

# Descriptive Inorganic Chemistry

Third Edition

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# Preface to the Third Edition

The present edition of Descriptive Inorganic Chemistry is based on the objectives that were described in the preface of the second edition. Early chapters provide a tool kit for understanding the structures and reactions that are so important in inorganic chemistry. Of necessity, a brief introduction is provided to the language and approaches of quantum mechanics. In order to provide a more logical separation of topics, Chapter 2 provides essential information on the structure and properties of atoms, and Chapter 3 presents the basic ideas of covalent bonding and symmetry. Following the discussion of structures of solids, emphasis is placed on molecular polarity and the importance of intermolecular interactions, which provide a basis for understanding physical properties of inorganic substances.

In succeeding chapters, the chemistry of elements is presented in an order based on the periodic table. In these chapters, material has been added in numerous places in order to present new information that is relevant and/or timely. Several of the newly presented topics deal with environmental issues. We believe that the result is a more balanced and significant coverage of the field.

In order to show the importance of inorganic chemistry to the entire field of chemistry, we have added Chapter 23, which presents a potpourri of topics that range from uses of iron compounds in treating anemia in oak trees to the use of auranofin, cisplatin, and chloroquine in medicine. The emphasis is placed on the essential factors related to structure and bonding from the standpoint of the inorganic constituents rather than biological functions. The latter are factors best left to courses in biology and biochemistry.

To provide a more appealing book, virtually all illustrations presented in the first two editions have been reconstructed. It must be emphasized that, though we are not graphic artists, we have produced all illustrations. If some of the results look somewhat amateurish, it is because this book is author illustrated rather than professionally illustrated. However, we believe that the illustrations are appropriate and convey the essential information.

It is our opinion that this book meets the objectives of including about as much inorganic chemistry as most students would assimilate in a one-semester course, that the material chosen is appropriate, and that the presentation is lucid and accurate. It is to be hoped that users of this book will agree. Perhaps Dr. Youmans said it best in 1854:

*Every experienced teacher understands the necessity of making the acquisition of the elementary and foundation principles upon which a science rests, the first business of study. If these are thoroughly mastered, subsequent progress is easy and certain.*

Edward L. Youmans, Chemical Atlas; or the Chemistry of Familiar Objects, D. Appleton & Co., New York, 1854.

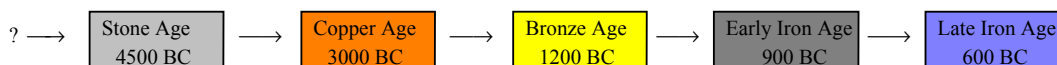
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## Chapter 1

# Where It All Comes From

Since the earliest times, man has sought for better materials to use in fabricating objects that were needed. Early man satisfied many requirements by gathering plants for food and fiber, and wood was used for making early tools and shelter. Stone and native metals, especially copper, were also used to make tools and weapons. The ages of man in history are generally identified by the materials that represented the dominant technology employed to fabricate useful objects. The approximate time periods corresponding to these epochs are designated as follows.



The biblical Old Testament period overlaps with the Copper, Bronze, and Iron Ages, so it is natural that these metals are mentioned frequently in the Bible and in other ancient manuscripts. For example, iron is mentioned about 100 times in the Old Testament, copper 8 times, and bronze more than 150 times. Other metals that were easily obtained (tin and lead) are also described numerous times. In fact, production of metals has been a significant factor in technology and chemistry for many centuries. Processes that are crude by modern standards were used many centuries ago to produce the desired metals and other materials, but the source of raw materials was the same then as it is now. In this chapter, we will present an overview of inorganic chemistry to show its importance in history and to relate it to modern industry.

### 1.1 THE STRUCTURE OF THE EARTH

There are approximately 16 million known chemical compounds, the vast majority of which are not found in nature. Although many of the known compounds are of little use or importance, some of them would be very difficult or almost impossible to live without. Try to visualize living in a world without concrete, synthetic fibers, fertilizer, steel, soap, glass, or plastics. None of these materials is found in nature in the form in which it is used, and yet they are all produced from naturally occurring raw materials. All of the items listed above and an enormous number of others are created by chemical processes. But created from *what?*

It has been stated that chemistry is the study of matter and its transformations. One of the major objectives of this book is to provide information on how the basic raw materials from the earth are transformed to produce inorganic compounds that are used on an enormous scale. It focuses attention on the transformations of a relatively few inorganic compounds available in nature into many others whether they are at present economically important or not. As you study this book, try to see the connection between obtaining a mineral by mining and the reactions that are used to convert it into end use products. Obviously, this book cannot provide the details for all such processes, but it does attempt to give an overview of inorganic chemistry and its methods and to show its relevance to the production of useful materials. Petroleum and coal are the major raw materials for organic compounds, but the transformation of these materials is not the subject of this book.

As it has been for all time, the earth is the source of all of the raw materials used in the production of chemical substances. The portion of the earth that is accessible for obtaining raw materials is that portion at the surface and slightly above and below the surface. This portion of the earth is referred to in geologic terms as the earth's crust. For thousands of years, man has exploited this region to gather stone, wood, water, and plants. In more modern times, many other chemical raw materials have been taken from the earth and metals have been removed on a huge scale. Although the techniques have changed, we are still limited in access to the resources of the atmosphere, water, and at most, a few miles of depth in the earth. It is the materials found in these regions of the earth that must serve as the starting materials for all of our chemical processes.

Because we are at present limited to the resources of the earth, it is important to understand the main features of its structure. Our knowledge of the structure of the earth has been developed by modern geoscience, and the gross features shown in Figure 1.1 are now generally accepted. The distances shown are approximate, and they vary somewhat from one geographical area to another.

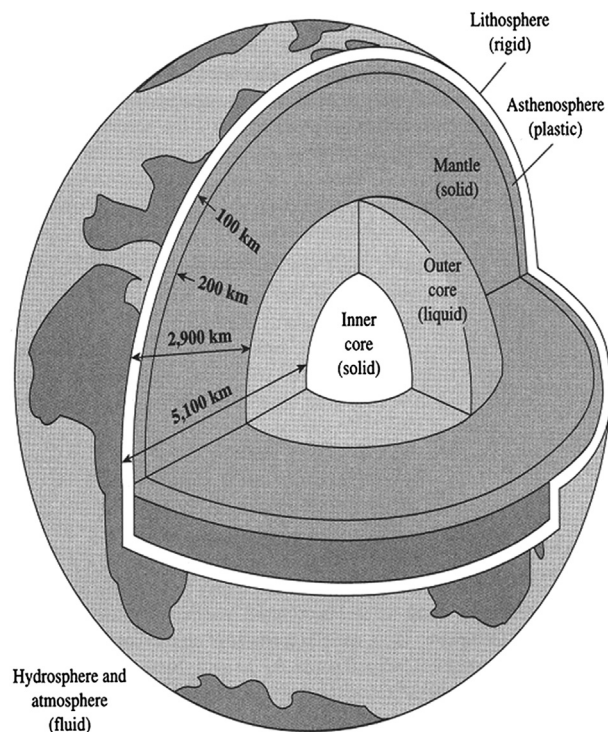


FIGURE 1.1 A cross section of the earth.

The region known as the *upper mantle* extends from the surface of the earth to a depth of approximately 660 km (400 mi). The *lower mantle* extends from a depth of about 660 km to about 3000 km (1800 mi). These layers consist of many substances, including some compounds that contain metals, but rocks composed of silicates are the dominant materials. The upper mantle is sometimes subdivided into the *lithosphere*, extending to a depth of approximately 100 km (60 mi), and the *asthenosphere*, extending from approximately 100 km to about 220 km (140 mi). The solid portion of the earth's crust is regarded as the lithosphere, and the *hydrosphere* and *atmosphere* are the liquid and gaseous regions, respectively. In the asthenosphere, the temperature and pressure are higher than in the lithosphere. As a result, it is generally believed that the asthenosphere is partially molten and softer than the lithosphere lying above it.

The core lies farther below the mantle, and two regions constitute the earth's core. The *outer core* extends from about 3000 km (1800 mi) to about 5000 km (3100 mi), and it consists primarily of molten iron. The *inner core* extends from about 5000 km to the center of the earth about 6500 km (4000 mi) below the surface, and it consists primarily of solid iron. It is generally believed that both core regions contain iron mixed with other metals, but iron is the major component.

The velocity of seismic waves shows unusual behavior in the region between the lower mantle and the outer core. The region where this occurs is at a much higher temperature than is the lower mantle, but it is cooler than the core. Therefore, the region has a large temperature gradient, and its chemistry is believed to be different from that of either the core or mantle. Chemical substances that are likely to be present include metallic oxides such as magnesium oxide and iron oxide, as well as silicon dioxide which is present as a form of *quartz* known as *stishovite* that is stable at high pressure. This is a region of very high pressure with estimates being as high as perhaps a million times that of the atmosphere. Under the conditions of high temperature and pressure, metal oxides react with  $\text{SiO}_2$  to form compounds such as  $\text{MgSiO}_3$  and  $\text{FeSiO}_3$ . Materials that are described by the formula  $(\text{Mg,Fe})\text{SiO}_3$  (where (Mg,Fe) indicates a material having a composition intermediate between the formulas above) are also produced.

## 1.2 COMPOSITION OF THE EARTH'S CRUST

Most of the elements shown in the periodic table are found in the earth's crust. A few have been produced artificially, but the rocks, minerals, atmosphere, lakes, and oceans have been the source of the majority of known elements. The abundance by mass of several elements that are major constituents in the earth's crust is shown in [Table 1.1](#).

**TABLE 1.1** Abundances of Elements by Mass

Element	O	Si	Al	Fe	Ca	Na	K	Mg	H	All Others
Percent	49.5	25.7	7.5	4.7	3.4	2.6	2.4	1.9	0.9	1.4

Elements such as chlorine, lead, copper, and sulfur occur in very small percentages, and although they are of great importance, they are relatively minor constituents. We must remember that there is a great difference between a material being *present*, and it being *recoverable* in a way that is *economically practical*. For instance, baseball-size nodules rich in manganese, iron, copper, nickel, and cobalt are found in large quantities on the ocean floor at a depth of 5–6 km. In addition, throughout the millennia, gold has been washed out of the earth and transported as minute particles to the oceans. However, it is important to understand that although the oceans are believed to contain vast quantities of metals including billions of tons of gold, there is at present no feasible way to recover these metals. Fortunately, compounds of some of the important elements are found in concentrated form in specific localities, and as a result they are readily accessible. It may be surprising to learn that even coal and petroleum that are used in enormous quantities are relatively minor constituents of the lithosphere. These complex mixtures of organic compounds are present to such a small extent that carbon is not among the most abundant elements. However, petroleum and coal are found concentrated in certain regions so they can be obtained by economically acceptable means. It would be quite different if all the coal and petroleum were distributed uniformly throughout the earth's crust.

### 1.3 ROCKS AND MINERALS

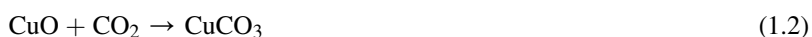
The chemical resources of early man were limited to the metals and compounds on the earth's surface. A few metals, e.g., copper, silver, and gold, were found uncombined (native) in nature so they have been available for many centuries. It is believed that the iron first used may have been found as uncombined iron that had reached the earth in the form of meteorites. In contrast, elements such as fluorine and sodium are produced by electrochemical reactions, and they have been available a much shorter time.

Most metals are found in the form of naturally occurring chemical compounds called *minerals*. An *ore* is a material that contains a sufficiently high concentration of a mineral to constitute an economically feasible source from which the metal can be recovered. Rocks are composed of solid materials that are found in the earth's crust, and they usually contain mixtures of minerals in varying proportions. Three categories are used to describe rocks based on their origin. Rocks that were formed by the solidification of a molten mass are called *igneous rocks*. Common examples of this type include *granite*, *feldspar*, and *quartz*. *Sedimentary rocks* are those which formed from compacting of small grains that have been deposited as a sediment in a river bed or sea, and they include such common materials as *sandstone*, *limestone*, and *dolomite*. Rocks that have had their composition and structure changed over time by the influences of temperature and pressure are called *metamorphic rocks*. Some common examples are *marble*, *slate*, and *gneiss*.

The lithosphere consists primarily of rocks and minerals. Some of the important classes of metal compounds found in the lithosphere are oxides, sulfides, silicates, phosphates, and carbonates. The atmosphere surrounding the earth contains oxygen so several metals such as iron, aluminum, tin, magnesium, and chromium are found in nature as the oxides. Sulfur is found in many places in the earth's crust (particularly in regions where there is volcanic activity) so some metals are found combined with sulfur as metal sulfides. Metals found as sulfides include copper, silver, nickel, mercury, zinc, and lead. A few metals, especially sodium, potassium, and magnesium, are found as the chlorides. Several carbonates and phosphates occur in the lithosphere, and calcium carbonate and calcium phosphate are particularly important minerals.

### 1.4 WEATHERING

Conditions on the inside of a rock may be considerably different from those at the surface. Carbon dioxide can be produced by the decay of organic matter, and an acid–base reaction between CO<sub>2</sub> and metal oxides produces metal carbonates. Typical reactions of this type are the following.



Moreover, because the carbonate ion can react as a base, it can remove  $\text{H}^+$  from water to produce hydroxide ions and bicarbonate ions by the following reaction.

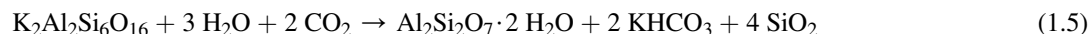


Therefore, as an oxide mineral “weathers,” reactions of  $\text{CO}_2$  and water at the surface lead to the formation of carbonates and bicarbonates. The presence of  $\text{OH}^-$  can eventually cause part of the mineral to be converted to a metal hydroxide. Because of the basicity of the oxide ion, most metal oxides react with water to produce hydroxides. An important example of such a reaction is



As a result of reactions such as these, a metal oxide may be converted by processes in nature to a metal carbonate or a metal hydroxide. A type of compound closely related to carbonates and hydroxides is known as a basic metal carbonate, and these materials contain both carbonate ( $\text{CO}_3^{2-}$ ) and hydroxide ( $\text{OH}^-$ ) ions. A well-known material of this type is  $\text{CuCO}_3 \cdot \text{Cu(OH)}_2$  or  $\text{Cu}_2\text{CO}_3(\text{OH})_2$  that is the copper-containing mineral known as *malachite*. Another mineral containing copper is *azurite* that has the formula  $2 \text{CuCO}_3 \cdot \text{Cu(OH)}_2$  or  $\text{Cu}_3(\text{CO}_3)_2(\text{OH})_2$  so it is quite similar to malachite. Azurite and malachite are frequently found together because both are secondary minerals produced by weathering processes. In both cases, the metal oxide,  $\text{CuO}$ , has been converted to a mixed carbonate/hydroxide compound. This example serves to illustrate how metals are sometimes found in compounds having unusual but closely related formulas. It also shows why ores of metals frequently contain two or more minerals containing the same metal.

Among the most common minerals are the feldspars and clays. These materials have been used for centuries in the manufacture of pottery, china, brick, cement, and other materials. Feldspars include the mineral *orthoclase*,  $\text{K}_2\text{O} \cdot \text{Al}_2\text{O}_3 \cdot 6 \text{SiO}_2$ , but this formula can also be written as  $\text{K}_2\text{Al}_2\text{Si}_6\text{O}_{16}$ . Under the influence of carbon dioxide and water, this mineral weathers by a reaction that can be shown as



The product,  $\text{Al}_2\text{Si}_2\text{O}_7 \cdot 2 \text{H}_2\text{O}$ , is known as *kaolinite* and it is one of the aluminosilicates that constitutes clays used in making pottery and china. This example also shows how one mineral can be converted into another by the natural process of weathering.

## 1.5 OBTAINING METALS

Because of their superior properties, metals have received a great deal of attention since the earliest times. Their immense importance now as well as throughout history indicates that we should describe briefly the processes involved in the production and use of metals. The first metal to be used extensively was copper because of its being found uncombined, but most metals are found combined with other elements in minerals. Minerals are naturally occurring compounds or mixtures of compounds that contain chemical elements. As we have mentioned, a mineral may *contain* some desired metal, but it may not be available in sufficient quantity and purity to serve as a useful *source* of the metal. A commercially usable source of a desired metal is known as an ore.

Most ores are obtained by mining. In some cases, ores are found on or near the surface making it possible for them to be obtained easily. In order to exploit an ore as a useful source of a metal, a large quantity of the ore is usually required. Two of the procedures still used today to obtain ores have been used for centuries. One of these methods is known as *open-pit mining*, and in this technique the ore is recovered by digging in the earth's surface. A second type of mining is *shaft mining* in which a shaft is dug into the earth to gain access to the ore below the surface. Coal and the ores of many metals are obtained by both of these methods. In some parts of the country, huge pits can be seen where the ores of copper and iron have been removed in enormous amounts. In other areas, the evidence of strip mining coal is clearly visible. Of course, the massive effects of shaft mining are much less visible.

Although mechanization makes mining possible on an enormous scale today, mining has been important for millennia. We know from ancient writings such as the Bible that mining and refining of metals have been carried for thousands of years (for example, see Job Chapter 28). Different types of ores are found at different depths, so both open-pit and shaft mining are still in common use. Coal is mined by both open-pit (strip mining) and shaft methods. Copper is mined by the open-pit method in Arizona, Utah, and Nevada, and iron is obtained in this way in Minnesota.

After the metal-bearing ore is obtained, the problem is to obtain the metal from the ore. Frequently, an ore may not have a high enough content of the mineral containing the metal to use it directly. The ore usually contains

varying amounts of other materials (rocks, dirt, etc.), which is known as *gangue* (pronounced “gang”). Before the mineral can be reduced to produce the free metal, the ore must be concentrated. Today, copper ores containing less than 1% copper are processed to obtain the metal. In early times, concentration consisted of simply picking out the pieces of the mineral by hand. For example, copper-containing minerals are green in color so they were easily identified. In many cases, the metal may be produced in a smelter located far from the mine. Therefore, concentrating the ore at the mine site saves on transportation costs and helps prevent the problems associated with disposing of the gangue at the smelting site.

The remaining gangue must be removed, and the metal must be reduced and purified. These steps constitute the procedures referred to as *extractive metallurgy*. After the metal is obtained, a number of processes may be used to alter its characteristics of hardness, workability, etc. The processes used to bring about changes in properties of a metal are known as *physical metallurgy*.

The process of obtaining metals from their ores by heating them with reducing agents is known as *smelting*. Smelting includes the processes of concentrating the ore, reducing the metal compound to obtain the metal, and purifying the metal. Most minerals are found mixed with a large amount of rocky material that usually is composed of silicates. In fact, the desired metal compound may be a relatively minor constituent in the ore. Therefore, before further steps to obtain the metal can be undertaken, the ore must be concentrated. Several different procedures are useful to concentrate ores depending on the metal.

The *flotation process* consists of grinding the ore to a powder and mixing it with water, oil, and detergents (wetting agents). The mixture is then beaten into a froth. The metal ore is concentrated in the froth so it can be skimmed off. For many metals, the ores are more dense than the silicate rocks, dirt and other material that contaminate them. In these cases, passing the crushed ore down an inclined trough with water causes the heavier particles of ore to be separated from the gangue.

*Magnetic separation* is possible in the case of the iron ore *taconite*. The major oxide in taconite is  $\text{Fe}_3\text{O}_4$  (this formula also represents  $\text{FeO} \cdot \text{Fe}_2\text{O}_3$ ) that is attracted to a magnet. The  $\text{Fe}_3\text{O}_4$  can be separated from most of the gangue by passing the crushed ore on a conveyor under a magnet. During the reduction process, removal of silicate impurities can also be accomplished by the addition of a material that forms a compound with them. When heated at high temperatures, *limestone*,  $\text{CaCO}_3$ , reacts with silicates to form a molten slag that has a lower density than the molten metal. The molten metal can be drained from the bottom of the furnace or the floating slag can be skimmed off the top.

After the ore is concentrated, the metal must be reduced from the compound containing it. Production of several metals will be discussed in later chapters of this book. However, a reduction process that has been used for thousands of years will be discussed briefly here. Several reduction techniques are now available, but the original procedure involved reduction of metals using carbon in the form of charcoal. When ores containing metal sulfides are heated in air (known as *roasting* the ore), they are converted to the metal oxides. In the case of copper sulfide, the reaction is



In recent years, the  $\text{SO}_2$  from this process has been trapped and converted into sulfuric acid. Copper oxide can be reduced using carbon as the reducing agent in a reaction that can be represented by the following equation.



For the reduction of  $\text{Fe}_2\text{O}_3$ , the equation can be written as



Because some metals are produced in enormous quantities, it is necessary that the reducing agent be readily available in large quantities and be inexpensive. Consequently, carbon is used as the reducing agent. When coal is heated strongly, volatile organic compounds are driven off and carbon is left in the form of *coke*. This is the reducing agent used in the production of several metals.

Extractive metallurgy today involves three types of processes. *Pyrometallurgy* refers to the use of high temperatures to bring about smelting and refining of metals. *Hydrometallurgy* refers to the separation of metal compounds from ores by the use of aqueous solutions. *Electrometallurgy* refers to the use of electricity to reduce the metal from its compounds.

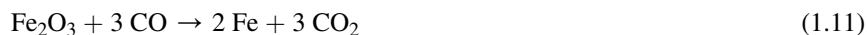
In ancient times, pyrometallurgy was used exclusively. Metal oxides were reduced by heating them with charcoal. The ore was broken into small pieces and heated in a stone furnace on a bed of charcoal. Remains of these ancient furnaces can still be observed in areas of the Middle East. Such smelting procedures are not very efficient, and the rocky material remaining after removal of the metal (known as *slag*) contained some unrecovered metal. Slag heaps from ancient smelting



furnaces show clearly that copper and iron smelting took place in the region of the Middle East known as the Arabah many centuries ago. Incomplete combustion of charcoal produces some carbon monoxide,



and carbon monoxide may also cause the reduction of some of the metal oxide as shown in these reactions.



Carbon monoxide is also an effective reducing agent in the production of metals today.

Because of its ease of reduction, copper was the earliest metal smelted. It is believed that the smelting of copper took place in the Middle East as early as about 2500–3500 BC. Before the reduction was carried out in furnaces, copper ores were probably heated in wood fires at a much earlier time. The metal produced in a fire or a crude furnace was impure so it had to be purified. Heating some metals to melting causes the remaining slag (called *dross*) to float on the molten metal where it can be skimmed off or the metal can be drained from the bottom of the melting pot. The melting process, known as *cupellation*, is carried out in a crucible or “fining” pot. Some iron refineries at Tel Jemmeh have been dated from about 1200 BC, the early iron age. The reduction of iron requires a higher temperature than that for the reduction of copper so smelting of iron occurred at a later time.

Although copper may have been used for perhaps 8000–10,000 years, the *reduction* of copper ores to produce the metal has been carried out since perhaps 4000 BC. The reduction of iron was practiced by about 1500–2000 BC (the Iron Age). Tin is easily reduced and somewhere in time between the use of charcoal to reduce copper and iron, the reduction of tin came to be known. Approximately 80 elements are metals and approximately 50 of them have some commercial importance. However, there are hundreds of *alloys* that have properties that make them extremely useful for certain applications. The development of alloys such as stainless steel, magnesium alloys, and *Duriron* (an alloy of iron and silicon) has occurred in modern times. Approximately 2500 BC it was discovered that adding about 3–4% of tin to copper made an alloy that has greatly differing properties from those of copper alone. That alloy, bronze, became one of the most important materials, and its widespread use resulted in the Bronze Age. Brass is an alloy of copper and zinc. Although brass was known several centuries BC, zinc was not known as an element until 1746. It is probable that minerals containing zinc were found along with those containing copper, and reduction of the copper also resulted in the reduction of zinc producing a mixture of the two metals. It is also possible that some unknown mineral was reduced to obtain an impure metal without knowing that the metal was zinc. Deliberately adding metallic zinc reduced from other sources to copper to make brass would have been unlikely because zinc was not a metal known in ancient times and it is more difficult to reduce than copper.

After a metal is obtained, there remains the problem of making useful objects from the metal, and there are several techniques that can be used to shape the object. In modern times, rolling, forging, spinning, and other techniques are used in fabricating objects from metals. In ancient times, one of the techniques used to shape metals was by hammering the cold metal. Hammered metal objects have been found in excavations throughout the world.

*Cold working* certain metals causes them to become harder and stronger. For example, if a wire made of iron is bent to make a kink in it, the wire will break at that point after flexing it a few times. When a wire made of copper is treated in this way, flexing it a few times causes the wire to bend in a new location beside the kink. The copper wire does not break, and this occurs because flexing the copper makes it harder and stronger. In other words, the metal has had its properties altered by cold working it.

When a hot metal is shaped or “worked” by *forging*, the metal retains its softer, more ductile original condition when it cools. In the hot metal, atoms have enough mobility to return to their original bonding arrangements. The metal can undergo great changes in shape without work hardening occurring, which might make it unsuitable for the purpose intended. Cold working by hammering and hot working (forging) of metal objects have been used in the fabrication of metal objects for many centuries.

## 1.6 SOME METALS TODAY

Today, as in ancient times, our source of raw materials is the earth’s crust. However, because of our advanced chemical technology, exotic materials have become necessary for processes that are vital yet unfamiliar to most people. This is true even for students in chemistry courses at the university level. For example, a chemistry student may know little about niobium or *bauxite*, but these materials are vital to our economy.

**TABLE 1.2** Some Inorganic Raw Materials

Material	Major Uses of Products	Sources	Percent Imported
Bauxite	Aluminum, abrasives, refractories, Al <sub>2</sub> O <sub>3</sub>	Brazil, Australia, Jamaica, Guyana	100
Niobium	Special steels, titanium alloys	Canada, Brazil	100
Graphite	Lubricants, crucibles, electrical components, pencils, nuclear moderator	Mexico, Canada, Sri Lanka, Madagascar	100
Manganese	Special steels, paints, batteries	S. Africa, Brazil, France, Australia	100
Mica	Electrical equipment, paints	India, Brazil, China, Belgium	100
Strontium	Glasses, ceramics, paints, TV tubes	Mexico	100
Rare earth metals	Permanent magnets for hybrid vehicles and wind turbines, phosphors for cell phones and computers, catalysts	China	~ 100
Diamonds	Cutting tools, abrasives	S. Africa, Zaire	98
Fluorite	HF, steel making	Mexico, Morocco, S. Africa, Canada	89
Platinum	Catalysts, alloys, metals (Pt, dental uses, Pd, Rh, Ir, surgical appliances Ru, Os)	S. Africa, Russia	88
Tantalum	Electronic capacitors, chemical equipment	Germany, Canada, Brazil, Australia	86
Chromium	Stainless steel, leather tanning, plating, alloys	S. Africa, Turkey, Zimbabwe	82
Tin	Alloys, plating, making flat glass	Bolivia, Brazil, Peru, Malaysia	81
Cobalt	Alloys, catalysts, magnets	Zambia, Zaire, China, Canada, Norway	75
Cadmium	Alloys, batteries, plating, reactors	Canada, Australia, Mexico	66
Nickel	Batteries, plating, coins, catalysts	Canada, Norway, Australia	64

An additional feature that makes obtaining many inorganic materials so difficult is that they are not distributed uniformly in the earth's crust. It is a fact of life that the major producers of niobium are Canada and Brazil, and the United States imports 100% of the niobium needed. The situation is similar for *bauxite*, major deposits of which are found in Brazil, Jamaica, Australia, and French Guyana. In fact, of the various ores and minerals that are sources of important inorganic materials, the United States must rely on other countries for many of them. Table 1.2 shows some of the major inorganic raw materials, their uses, and their sources.

The information shown in Table 1.2 reveals that no industrialized country is entirely self-sufficient in terms of all necessary natural resources. In many cases, metals are recycled, so that the need for imports is lessened. For instance, although 100% of the ore bauxite is imported, approximately 37% of the aluminum used in the United States comes from recycling. In 2014, 50,000 kg of platinum was recovered from catalytic converters. About 36% of the chromium, 41% of nickel, and 27% of the cobalt used are recovered from recycling. Changing political regimes may result in shortages of critical materials. In the 1990s, inexpensive imports of rare earth metals from China forced the closure of mines in the United States. Because of rising costs and the increased demand for rare earth metals in high performance batteries, a mine at Mountain Pass, California opened in 2014. Although the data shown in Table 1.2 paint a rather bleak picture of our metal resources, the United States is much better supplied with many nonmetallic raw materials.

## 1.7 NONMETALLIC INORGANIC MINERALS

Many of the materials that are so familiar to us are derived from petroleum or other organic sources. This is also true for the important polymers and an enormous number of organic compounds that are derived from organic raw materials. Because of the content of this book, we will not deal with this vast area of chemistry, but rather will discuss inorganic materials and their sources.

In ancient times, the chemical operations of reducing metals ores, making soap, dyeing fabric, etc., were carried out in close proximity to where people lived. These processes were familiar to most people of that day. Today, mines and factories may be located in remote areas or they may be separated from residential areas so that people have no knowledge of where the items come from or how they are produced. As chemical technology has become more sophisticated, a smaller percentage of people understand its operation and scope.

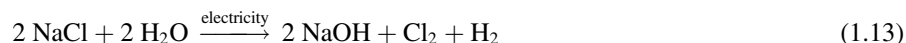
A large number of inorganic materials are found in nature. The chemical compound used in the largest quantity is sulfuric acid,  $\text{H}_2\text{SO}_4$ . It is arguably the most important single compound, and although approximately 79 billion pounds are used annually in the United States, it is not found in nature. However, sulfur is found in nature, and it is burned to produce sulfur dioxide that is oxidized in the presence of platinum as a catalyst to give  $\text{SO}_3$ . When added to water,  $\text{SO}_3$  reacts to give  $\text{H}_2\text{SO}_4$ . Also found in nature are metal sulfides. When these compounds are heated in air, they are converted to metal oxides and  $\text{SO}_2$ . The  $\text{SO}_2$  is utilized to make sulfuric acid, but the process described requires platinum (from Russia or South Africa) for use as a catalyst.

Another chemical used in large quantities (about 42 billion pounds annually) is lime,  $\text{CaO}$ . Like sulfuric acid, it is not found in nature, but it is produced from calcium carbonate which is found in several forms in many parts of the world. The reaction by which lime has been produced for thousands of years is



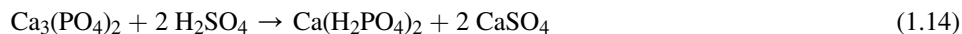
Lime is used in making glass, cement, and many other materials. Cement is used in making concrete, the material used in the largest quantity of all. Glass is not only an important material for making food containers, but also an extremely important construction material.

Salt is a naturally occurring inorganic compound. Although salt is of considerable importance in its own right, it is also used to make other inorganic compounds. For example, the electrolysis of an aqueous solution of sodium chloride produces sodium hydroxide, chlorine, and hydrogen.



Both sodium hydroxide and chlorine are used in the preparation of an enormous number of materials, both inorganic and organic.

Calcium phosphate is found in many places in the earth's crust. It is difficult to overemphasize its importance because it is used on an enormous scale in the manufacture of fertilizers by the reaction

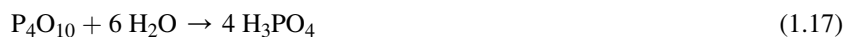


The  $\text{Ca}(\text{H}_2\text{PO}_4)_2$  is preferable to  $\text{Ca}_3(\text{PO}_4)_2$  for use as a fertilizer because it is more soluble in water. The  $\text{CaSO}_4$  is known as gypsum and, although natural gypsum is mined in some places, that produced by the reaction above is an important constituent in wallboard. The reaction above is carried out on a scale that is almost unbelievable. About 50% of the over 79 billion pounds of  $\text{H}_2\text{SO}_4$  used annually in the United States goes into the production of fertilizers. With a world population that has reached 7 billion, the requirement for foodstuffs would be impossible to meet without effective fertilizers.

Calcium phosphate is an important raw material in another connection. It serves as the source of elemental phosphorus that is produced by the following reaction.



Phosphorus reacts with chlorine to yield  $\text{PCl}_3$  and  $\text{PCl}_5$ . These are reactive substances that serve as the starting materials for making many other materials that contain phosphorus. Moreover,  $\text{P}_4$  burns in air to yield  $\text{P}_4\text{O}_{10}$  which reacts with water to produce phosphoric acid, another important chemical of commerce, as shown in the following equations.



Only a few inorganic raw materials have been mentioned and their importance described very briefly. The point of this discussion is to show that although a large number of inorganic chemicals are useful, they are not found in nature in the forms needed. It is the *transformation* of raw materials into the many other useful compounds that is the subject of this book. As you study this book, keep in mind that the processes shown are relevant to the production of inorganic compounds that are vital to our way of life.

**TABLE 1.3** Important Inorganic Chemicals

Compound	2014 Production, Billion lbs	Uses
H <sub>2</sub> SO <sub>4</sub>	79	Fertilizers, chemicals, batteries
N <sub>2</sub>	75	Fertilizers
O <sub>2</sub>	61	Steel production, welding
Lime, CaO	42	Metals reduction, chemicals, water treatment
NH <sub>3</sub>	18	Fertilizers, polymers, explosives
H <sub>3</sub> PO <sub>4</sub>	22	Fertilizers, chemicals, foods
Cl <sub>2</sub>	24	Bleaches, chemicals, water treatment
Sulfur	24	Sulfuric acid, detergents, chemicals
Na <sub>2</sub> CO <sub>3</sub>	24	Glass, chemicals, laundry products
NaOH	31	Chemicals, paper, soaps
HNO <sub>3</sub>	19	Fertilizers, explosives, propellants
Urea <sup>a</sup>	13	Fertilizers, animal feeds, polymers
NH <sub>4</sub> NO <sub>3</sub>	14	Fertilizers, explosives
HCl	9.7	Metal treatment, chemicals

<sup>a</sup>An "organic" compound produced by the reaction of NH<sub>3</sub> and CO<sub>2</sub>.

In addition to the inorganic raw materials shown in Table 1.2, a very brief mention has been made of a few of the most important inorganic chemicals. Although many other inorganic compounds are needed, Table 1.3 shows some of the inorganic compounds that are produced in the largest quantities in the United States. Of these, only N<sub>2</sub>, O<sub>2</sub>, sulfur, and Na<sub>2</sub>CO<sub>3</sub> occur naturally. Many of these materials will be discussed in later chapters, and in some ways they form the core of industrial inorganic chemistry. As you study this book, note how frequently the chemicals listed in Table 1.3 are mentioned and how processes involving them are of such great economic importance.

As you read this book, also keep in mind that it is not possible to remove natural resources without producing some environmental changes. Certainly, every effort should be made to lessen the impact of all types of mining operations on the environment and landscape. Steps must also be taken to minimize the impact of chemical industries on the environment. However, as we drive past a huge hole where open-pit mining of iron ore has been carried out, we must never forget that without the ore being removed there would be nothing to drive.

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## PROBLEMS

1. What are the names of the solid, liquid, and gaseous regions of the earth's crust?
2. What metal is the primary component of the earth's core?
3. Elements such as copper and silver are present in the earth's crust in very small percentages. What is it about these elements that makes their recovery economically feasible?
4. Explain the difference between rocks, minerals, and ores.

5. How were igneous rocks such as granite and quartz formed?
6. How were sedimentary rocks such as limestone and dolomite formed?
7. How were metamorphic rocks such as marble and slate formed?
8. What are some of the important classes of metal compounds found in the lithosphere?
9. Write the chemical equations that show how the process of weathering leads to formation of carbonates and hydroxides.
10. Why was copper the first metal to be used extensively?
11. Describe the two types of mining used to obtain ores.
12. Describe the procedures used to concentrate ores.
13. Metals are produced in enormous quantities. What two properties must a reducing agent have in order to be used in the commercial refining of metals?
14. Describe the three types of processes used in extractive metallurgy.
15. What was the earliest metal smelted? Why was iron not smelted until a later time?
16. Name three modern techniques used to shape metals.
17. Name two ancient techniques used to shape metals.
18. Briefly describe what the effect on manufacturing might be if the United States imposed a total trade embargo on a country such as South Africa.
19. Approximately 81 billion pounds of sulfuric acid are used annually. What inorganic material is the starting material in the manufacture of sulfuric acid?
20. What are some of the primary uses for lime,  $\text{CaO}$ ?
21. What is the raw material calcium phosphate,  $\text{Ca}_3(\text{PO}_4)_2$ , used primarily for?

## Chapter 2

# Atomic Structure and Properties

The fundamental unit involved in elements and the formation of compounds is the atom. Properties of atoms such as the energy necessary to remove an electron (ionization potential), energy of attraction for additional electrons (electron affinity), and atomic sizes are important factors that determine the chemical behavior of elements. Also, the arrangement of electrons in atoms has a great deal of influence on the types of molecules the atoms can form. Because descriptive inorganic chemistry is the study of chemical reactions and properties of molecules, it is appropriate to begin that study by presenting an overview of the essentials of atomic structure.

The structure of atoms is based on the fundamental principles described in courses such as atomic physics and quantum mechanics. In a book of this type, it is not possible to present more than a cursory description of the results obtained by experimental and theoretical studies on atomic structure. Consequently, what follows is a nonmathematical treatment of the aspects of atomic structure that provides an adequate basis for understanding much of the chemistry presented later in this book. Much of this chapter should be a review of principles learned in earlier chemistry courses, which is intentional. More theoretical treatments of these topics can be found in the suggested readings at the end of this chapter.

## 2.1 ATOMIC STRUCTURE

A knowledge of the structure of atoms provides the basis for understanding how they combine and the type of bonds that are formed. In this section, a review of early work in this area will be presented and variations in atomic properties will be related to the periodic table.

### 2.1.1 Quantum Numbers

It was the analysis of the line spectrum of hydrogen observed by J. J. Balmer and others that led Niels Bohr to a treatment of the hydrogen atom that is now referred to as the *Bohr model*. In that model, there are supposedly “allowed” orbits in which the electron can move around the nucleus without radiating electromagnetic energy. The orbits are those for which the angular momentum,  $mvr$ , can have only certain values (they are referred to as being *quantized*). This condition can be represented by the relationship

$$mvr = \frac{nh}{2\pi} \quad (2.1)$$

where  $n$  is an integer (1, 2, 3, ...) corresponding to the orbit,  $h$  is *Planck's constant*,  $m$  is the mass of the electron,  $v$  is its velocity, and  $r$  is the radius of the orbit. Although the Bohr model gave a successful interpretation of the line spectrum of hydrogen, it did not explain the spectral properties of species other than hydrogen and ions containing a single electron ( $\text{He}^+$ ,  $\text{Li}^{2+}$ , etc.).

In 1924, Louis de Broglie, as a young doctoral student, investigated some of the consequences of relativity theory. It was known that for electromagnetic radiation, the energy,  $E$ , is expressed by the Planck relationship,

$$E = hv = \frac{hc}{\lambda} \quad (2.2)$$

where  $c$ ,  $v$ , and  $\lambda$  are the velocity, frequency, and wavelength of the radiation, respectively. The photon also has an energy given by a relationship obtained from relativity theory,

$$E = mc^2 \quad (2.3)$$

A specific photon can have only one energy so the right-hand sides of [Eqs \(2.2\) and \(2.3\)](#) must be equal. Therefore,

$$\frac{hc}{\lambda} = mc^2 \quad (2.4)$$

and solving for the wavelength gives

$$\lambda = \frac{h}{mc} \quad (2.5)$$

The product of mass and velocity equals momentum so the wavelength of a photon, represented by  $h/mc$ , is Planck's constant divided by its momentum. Because particles have many of the characteristics of photons, de Broglie reasoned that for a *particle* moving at a velocity,  $v$ , there should be an associated wavelength that is expressed as

$$\lambda = \frac{h}{mv} \quad (2.6)$$

This predicted wave character was verified in 1927 by C. J. Davisson and L. H. Germer who studied the diffraction of an electron beam that was directed at a nickel crystal. Diffraction is a characteristic of waves so it was demonstrated that moving electrons have a wave character.

If an electron behaves as a *wave* as it moves in a hydrogen atom, a stable orbit can result only when the circumference of a circular orbit contains a whole number of waves. In that way, the waves can join smoothly to produce a standing wave with the circumference being equal to an integral number of wavelengths. This equality can be represented as

$$2\pi r = n\lambda \quad (2.7)$$

where  $n$  is an integer. Because  $\lambda$  is equal to  $h/mv$ , substitution of this value in Eq. (2.7) gives

$$2\pi r = n \frac{h}{mv} \quad (2.8)$$

which can be rearranged to give

$$mvr = \frac{nh}{2\pi} \quad (2.9)$$

It should be noted that this relationship is identical to Bohr's *assumption* about stable orbits (shown in Eq. (2.1))!

In 1926, Erwin Schrödinger made use of the wave character of the electron and adapted a previously known equation for three-dimensional waves to the hydrogen atom problem. The result is known as the Schrödinger wave equation for the hydrogen atom which can be written as

$$\nabla^2 \Psi + \frac{2m}{\hbar^2} (E - V) \Psi = 0 \quad (2.10)$$

where  $\Psi$  is the *wave function*,  $\hbar$  is  $h/2\pi$ ,  $m$  is the mass of the electron,  $E$  is the total energy,  $V$  is the potential energy (in this case the electrostatic energy) of the system, and  $\nabla^2$  is the Laplacian operator.

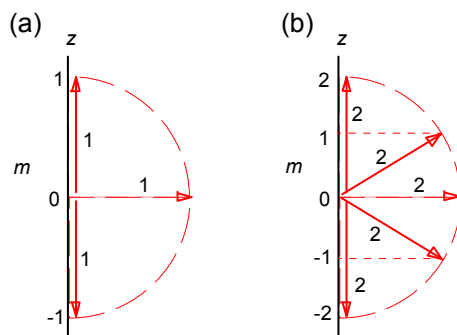
$$\nabla^2 = \frac{\partial^2}{\partial x^2} + \frac{\partial^2}{\partial y^2} + \frac{\partial^2}{\partial z^2} \quad (2.11)$$

The wave function is, therefore, a function of the coordinates of the parts of the system that completely describes the system. A useful characteristic of the quantum mechanical way of treating problems is that once the wave function is known, it provides a way for calculating some properties of the system.

The Schrödinger equation for the hydrogen atom is a second-order partial differential equation in three variables. A customary technique for solving this type of differential equation is by a procedure known as the separation of variables. In that way, a complicated equation that contains multiple variables is reduced to multiple equations, each of which contains a smaller number of variables. The potential energy,  $V$ , is a function of the distance of the electron from the nucleus, and this distance is represented in Cartesian coordinates as  $r = (x^2 + y^2 + z^2)^{1/2}$ . Because of this relationship, it is impossible to use the separation of variables technique. Schrödinger solved the wave equation by first transforming the Laplacian operator into polar coordinates. The resulting equation can be written as

$$\frac{1}{r^2} \frac{\partial}{\partial r} r^2 \frac{\partial \Psi}{\partial r} + \frac{1}{r^2 \sin \theta} \frac{\partial}{\partial \theta} \left( \sin \theta \frac{\partial \Psi}{\partial \theta} \right) + \frac{1}{r^2 \sin^2 \theta} \frac{\partial^2 \Psi}{\partial \phi^2} + \frac{2m}{\hbar^2} \left( E + \frac{e^2}{r} \right) \Psi = 0 \quad (2.12)$$

Although no attempt will be made to solve this very complicated equation, it should be pointed out that in this form the separation of the variables is possible, and equations that are functions of  $r$ ,  $\theta$ , and  $\phi$  result. Each of the simpler equations that are obtained can be solved to give solutions that are functions of only one variable. These partial solutions are



**FIGURE 2.1** Illustrations of the possible  $m_l$  values for cases where  $l = 1$  (a) and  $l = 2$  (b).

described by the functions  $R(r)$ ,  $\Theta(\theta)$ , and  $\Phi(\phi)$ , respectively, and the overall solution is the product of these partial solutions.

It is important to note at this point that the mathematical restrictions imposed by solving the differential equations naturally lead to some restraints on the nature of the solutions. For example, solution of the equation containing  $r$  requires the introduction of an integer,  $n$ , which can have the values  $n = 1, 2, 3, \dots$  and an integer  $l$ , which has values that are related to the value of  $n$  such that  $l = 0, 1, 2, \dots (n - 1)$ . For a given value of  $n$ , the values for  $l$  can be all integers from 0 up to  $(n - 1)$ . The quantum number  $n$  is called the *principal quantum number* and  $l$  is called the *angular momentum quantum number*. The principal quantum number determines the energy of the state for the hydrogen atom but for complex atoms the energy also depends on  $l$ .

The partial solution of the equation that contains the angular dependence results in the introduction of another quantum number,  $m_l$ . This number is called the *magnetic quantum number*. The magnetic quantum number gives the quantized lengths of the projection of the  $l$  vector along the  $z$ -axis. Thus, this quantum number can take on values  $+l, (l - 1), \dots, 0, \dots, -l$ . This relationship is illustrated in Figure 2.1 for cases where  $l = 1$  and  $l = 2$ . If the atom is placed in a magnetic field, each of these states will represent a different energy. This is the basis for the *Zeeman effect*. One additional quantum number is required for a complete description of an electron in an atom because the electron has an intrinsic spin. The fourth quantum number is  $m_s$ , the *spin quantum number*. It is assigned values of  $+1/2$  or  $-1/2$  in units of  $h/2\pi$ , the quantum of angular momentum. Thus, a total of four quantum numbers ( $n, l, m_l$ , and  $m_s$ ) are required to completely describe an electron in an atom.

An energy state for an electron in an atom is denoted by writing the numerical value of the principal quantum number followed by a letter to denote the  $l$  value. The letters used to designate the  $l$  values 0, 1, 2, 3, ... are  $s, p, d, f, \dots$  respectively. These letters have their origin in the spectroscopic terms *sharp, principal, diffuse, and fundamental*, which are descriptions of the appearance of certain spectral lines. After the letter  $f$ , the sequence is alphabetical, except the letter  $j$  is not used. Consequently, states are denoted as  $1s, 2p, 3d, 4f$ , etc. There are no states such as  $1p, 2d$ , or  $3f$  because of the restriction that  $n \geq (l + 1)$ . Because  $l = 1$  for a  $p$  state, there will be three  $m_l$  values (0, +1, and -1) that correspond to three orbitals. For  $l = 2$  (corresponding to a  $d$  state), there are five values (+2, +1, 0, -1, and -2) possible for  $m_l$  so there are five orbitals in the  $d$  state.

## 2.1.2 Hydrogen-Like Orbitals

The wave functions for  $s$  states are functions of  $r$  and do not show any dependence on angular coordinates. Therefore, the orbitals represented by the wave functions are spherically symmetric, and the probability of finding the electron at a given distance from the nucleus in such an orbital is equal in all directions. This results in an orbital that can be shown as a spherical surface. Figure 2.2 shows an  $s$  orbital that is drawn to encompass the region where the electron will be found some fraction (perhaps 95%) of the time.

For  $p, d$ , and  $f$  states, the wave functions are mathematical expressions that contain a dependence on both distance ( $r$ ) and the coordinate angles  $\theta$  and  $\phi$ . As a result, these orbitals have directional character. A higher probability exists that the electron will be found in those regions, and the shapes of the regions of higher probability are shown in Figure 2.3 for  $p$  and  $d$  states. *The signs are the algebraic sign of the wave function in that region of space, not charges.*

The wave mechanical treatment of the hydrogen atom does not provide more accurate values than the Bohr model did for the energy states of the hydrogen atom. It does, however, provide the basis for describing the *probability* of finding electrons in certain regions, which is more compatible with the *Heisenberg uncertainty principle*. Note that the solution of



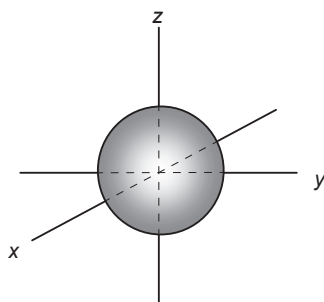


FIGURE 2.2 A spherical  $s$  orbital.

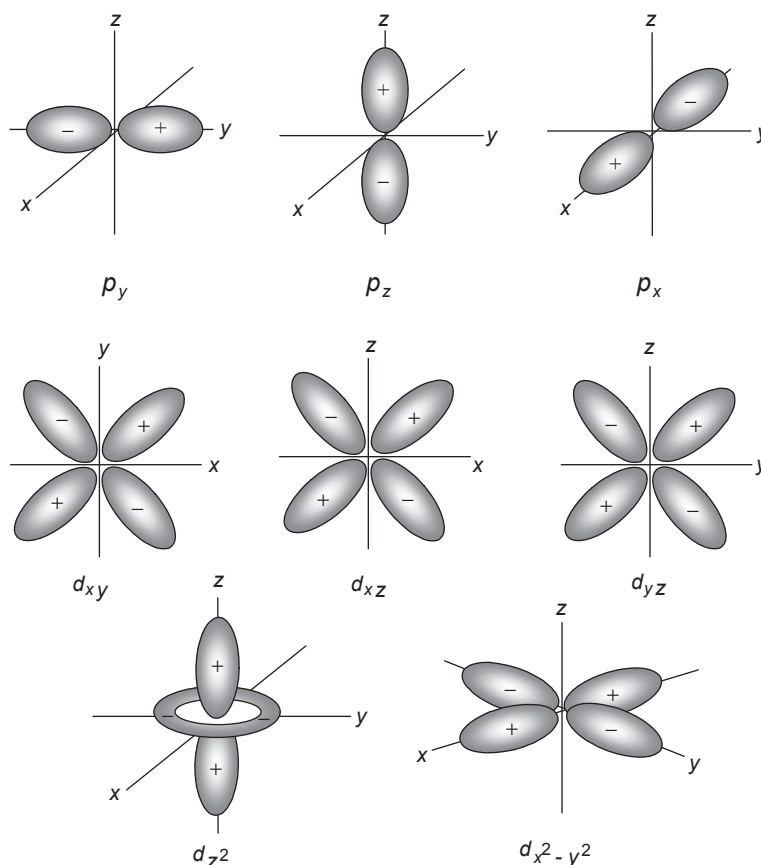


FIGURE 2.3 The three  $p$  orbitals and five  $d$  orbitals. The signs shown are the mathematical signs of the wave functions in the various regions of space. For ease of illustration, the orbital lobes are shown as ellipses rather than the actual shapes. This practice is followed in many places throughout this book.

this three-dimensional wave equation resulted in the introduction of three quantum numbers ( $n$ ,  $l$ , and  $m_l$ ). A principle of quantum mechanics predicts that there will be one quantum number for each dimension of the system being described by the wave equation. For the hydrogen atom, the Bohr model introduced only one quantum number,  $n$ , and that by an assumption.

## 2.2 PROPERTIES OF ATOMS

Although the solution of the wave equation has not been shown, it is still possible to make use of certain characteristics of the solutions. What is required is a knowledge of the properties of atoms. At this point, some of the empirical and experimental properties of atoms that are important for understanding descriptive chemistry will be described.

## 2.2.1 Electron Configurations

As has been mentioned, four quantum numbers are required to completely describe an electron in an atom, but there are certain restrictions on the values that these quantum numbers can have. For instance,  $n = 1, 2, 3, \dots$  and  $l = 0, 1, 2, \dots, (n - 1)$ . That is to say, for a given value of  $n$ , the quantum number  $l$  can have all integer values from 0 to  $(n - 1)$ . The quantum number  $m_l$  can have the series of values  $+l, +(l - 1), \dots, 0, \dots, -(l - 1), -l$ , so that there are  $(2l + 1)$  values for  $m_l$ . The fourth quantum number,  $m_s$  can have values of  $+\frac{1}{2}$  or  $-\frac{1}{2}$ , which is the spin angular momentum in units of  $h/2\pi$ . By making use of these restrictions, sets of quantum numbers can be written to describe electrons in atoms.

A necessary condition to be used is the *Pauli exclusion principle* which states that *no two electrons in the same atom can have the same set of four quantum numbers*. It should also be recognized that lower  $n$  values represent states of lower energy. For hydrogen, the four quantum numbers used to describe the single electron can be written as  $n = 1, l = 0, m_l = 0, m_s = +\frac{1}{2}$ . For convenience, the positive values of  $m_l$  and  $m_s$  are used before the negative values. For the two electrons in a helium atom the quantum numbers are as follows.

$$\begin{aligned} \text{Electron 1: } n = 1, l = 0, m_l = 0, m_s = +\frac{1}{2} \\ \text{Electron 2: } n = 1, l = 0, m_l = 0, m_s = -\frac{1}{2} \end{aligned}$$

Because an atomic energy level can be denoted by the  $n$  value followed by a letter ( $s, p, d$ , or  $f$  to denote  $l = 0, 1, 2$ , or  $3$ , respectively), the ground state for hydrogen is  $1s^1$  whereas that for helium is  $1s^2$ . The two sets of quantum numbers written above complete the first shell for which  $n = 1$ , and no other sets of quantum numbers are possible that have  $n = 1$ .

For  $n = 2$ ,  $l$  can have the values of 0 and 1. As a general rule, the levels increase in energy as the sum of  $n + l$  increases. Taking the value of  $l = 0$  first, the sets of quantum numbers are as follows.

$$\begin{aligned} \text{Electron 1: } n = 2, l = 0, m_l = 0, m_s = +\frac{1}{2} \\ \text{Electron 2: } n = 2, l = 0, m_l = 0, m_s = -\frac{1}{2} \end{aligned}$$

These two sets of quantum numbers describe electrons residing in the  $2s$  level. Taking next the  $l = 1$  value, it is found that six sets of quantum numbers can be written.

$$\begin{aligned} \text{Electron 1: } n = 2, l = 1, m_l = +1, m_s = +\frac{1}{2} \\ \text{Electron 2: } n = 2, l = 1, m_l = 0, m_s = +\frac{1}{2} \\ \text{Electron 3: } n = 2, l = 1, m_l = -1, m_s = +\frac{1}{2} \\ \text{Electron 4: } n = 2, l = 1, m_l = +1, m_s = -\frac{1}{2} \\ \text{Electron 5: } n = 2, l = 1, m_l = 0, m_s = -\frac{1}{2} \\ \text{Electron 6: } n = 2, l = 1, m_l = -1, m_s = -\frac{1}{2} \end{aligned}$$

These six sets of quantum numbers correspond to three pairs of electrons residing in the  $2p$  level. There are always as many orbitals as there are  $m_l$  values, each orbital capable of holding a pair of electrons, but the electrons remain unpaired as long as possible. For  $l = 2$  (which corresponds to a  $d$  state), there are five values for  $m_l$  ( $+2, +1, 0, -1$ , and  $-2$ ) and each can be used with  $m_s$  values of  $+\frac{1}{2}$  and  $-\frac{1}{2}$  so that a  $d$  state can hold 10 electrons. For an increase of 1 in the value of  $l$ , we gain two additional  $m_l$  values to which we can assign two values of  $m_s$ . Thus, there are always four more electrons possible for each successive state as shown in Table 2.1.

Except for minor variations that will be noted, the order of increasing energy levels in an atom is given by the sum  $(n + l)$ . The lowest value for  $(n + l)$  occurs when  $n = 1$  and  $l = 0$ , which corresponds to the  $1s$  state. The next lowest

**TABLE 2.1** Maximum Occupancy of Various Electron Shells

$l$ Value	$m_l$ Values	State	Maximum Number of Electrons
0	0	$s$	2
1	0, $\pm 1$	$p$	6
2	0, $\pm 1, \pm 2$	$d$	10
3	0, $\pm 1, \pm 2, \pm 3$	$f$	14
4	0, $\pm 1, \pm 2, \pm 3, \pm 4$	$g$	18

**TABLE 2.2** Energy States According to Increasing  $(n + l)$ 

$n$	$l$	$(n + l)$	State <sup>a</sup>	
1	0	1	1s	↓ I n c r e a s i n g  E
2	0	2	2s	
2	1	3	2p	
3	0	3	3s	
3	1	4	3p	
4	0	4	4s	
3	2	5	3d	
4	1	5	4p	
5	0	5	5s	
4	2	6	4d	
5	1	6	5p	
6	0	6	6s	
4	3	7	4f	
5	2	7	5d	
6	1	7	6p	
7	0	7	7s	

<sup>a</sup>It should be noted that this order is approximate and that the difference between successive states gets smaller farther down in the table. Thus, some irregularities in filling shells do occur.

sum of  $(n + l)$  is 2 when  $n = 2$  and  $l = 0$  (there is no  $1p$  state where  $n = 1$  and  $l = 1$  because  $l$  cannot equal  $n$ ). Continuing this process, we come to  $(n + l) = 4$ , which arises for  $n = 3$  and  $l = 1$  or  $n = 4$  and  $l = 0$ . Although the sum  $(n + l)$  is the same in both cases, the level with  $n = 3$  (the  $3p$  level) is filled first. When two or more ways exist for the same  $(n + l)$  sum to arise, the level with lower  $n$  will usually fill first. Table 2.2 shows the approximate order of filling the energy states.

Electron configurations of atoms can now be written by making use of the maximum occupancy and the order of filling the orbitals. The state of lowest energy is the ground state, and the electron configurations for all elements are shown in Appendix A. The filling of the states of lowest energy available is regular until Cr is reached. Here the configuration  $3d^4 4s^2$  is predicted, but it is  $3d^5 4s^1$  instead. The reason for this is the more favorable coupling of spin and orbital angular momenta that results when a greater number of unpaired electron spins interact, as is the case for a half-filled  $3d$  level. Therefore, for Cr, the configuration  $3d^5 4s^1$  represents a lower energy than does  $3d^4 4s^2$ . In the case of Cu, the electron configuration is  $3d^{10} 4s^1$  rather than  $3d^9 4s^2$  for the same reason.

The order of filling shells with electrons and the number of electrons that each shell can hold is reflected in the periodic table shown in Figure 2.4. Groups IA and IIA represent the groups where an  $s$  level is being filled as the outer shell whereas in Groups IIIA through VIIIA  $p$  shells fill in going from left to right. These groups where  $s$  or  $p$  levels are the outside shells are called the *main group elements*. First, second, and third series of transition elements are the rows where the  $3d$ ,  $4d$ , and  $5d$  levels are being filled. As a result, the elements in these groups are frequently referred to as “ $d$ -group elements.” Finally, the lanthanides and the actinides represent groups of elements where the  $4f$  and  $5f$  levels, respectively, are being filled.

The electron configurations and the periodic table show the similarities of electronic properties of elements in the same group. For example, the alkali metals (Group IA) all have an outside electronic arrangement of  $ns^1$ . As a result of the chemical properties of elements being strongly dependent on their outer (valence) shell electrons, it is apparent why elements in this group have so many chemical similarities. The halogens (Group VIIA) all have valence shell configurations of  $ns^2 np^5$ . Gaining an electron converts each to the configuration of the next noble gas,  $ns^2 np^6$ . It should be emphasized, however, that although there are many similarities, numerous differences also exist for elements in the same group. Thus, it should not be inferred that a similar electronic configuration in the valence shell gives rise to the same

IA 1												VIIIA 18					
1 H 1.0079	IIA 2											13 B 10.81	14 C 12.011	15 N 14.0067	16 O 15.9994	17 F 18.9984	18 Ne 20.179
3 Li 6.941	4 Be 9.0122											5 B 10.81	6 C 12.011	7 N 14.0067	8 O 15.9994	9 F 18.9984	10 Ne 20.179
11 Na 22.9898	12 Mg 24.305	III B 3	IV B 4	V B 5	V I B 6	V II B 7	VIII B 8 9 10		IB 11	IIB 12	13 Al 26.9815	14 Si 28.0855	15 P 30.9738	16 S 32.06	17 Cl 35.453	18 Ar 39.948	
19 K 39.0983	20 Ca 40.08	21 Sc 44.9559	22 Ti 47.88	23 V 50.9415	24 Cr 51.996	25 Mn 54.9380	26 Fe 55.847	27 Co 58.9332	28 Ni 58.69	29 Cu 63.546	30 Zn 65.38	31 Ga 69.72	32 Ge 72.59	33 As 74.9216	34 Se 78.96	35 Br 79.904	36 Kr 83.80
37 Rb 85.4678	38 Sr 87.62	39 Y 88.9059	40 Zr 91.22	41 Nb 92.9064	42 Mo 95.94	43 Tc (98)	44 Ru 101.07	45 Rh 102.906	46 Pd 106.42	47 Ag 107.868	48 Cd 112.41	49 In 114.82	50 Sn 118.69	51 Sb 121.75	52 Te 127.60	53 I 126.905	54 Xe 131.29
55 Cs 132.905	56 Ba 137.33	57 La* 138.906	72 Hf 178.48	73 Ta 180.948	74 W 183.85	75 Re 186.207	76 Os 190.2	77 Ir 192.22	78 Pt 195.09	79 Au 196.967	80 Hg 200.59	81 Tl 204.383	82 Pb 207.2	83 Bi 208.980	84 Po (209)	85 At (210)	86 Rn (222)
87 Fr (223)	88 Ra 226.025	89 Ac* 227.028	104 Rf (257)	105 Ha (260)	106 Sg (263)	107 Ns (262)	108 Hs (265)	109 Mt (266)	110 Ds (271)	111 Rg (272)	112 Cp* (285)	113 Uut (284)	114 Uuq (289)	115 Uup (288)	116 Uuh (293)	117 Uus (?)	118 Uuo (294)

*Lanthanide Series	58 Ce 140.12	59 Pr 140.908	60 Nd 144.24	61 Pm (145)	62 Sm 150.36	63 Eu 151.96	64 Gd 157.25	65 Tb 158.925	66 Dy 162.50	67 Ho 164.930	68 Er 167.26	69 Tm 168.934	70 Yb 173.04	71 Lu 174.967
*Actinide Series	90 Th 232.038	91 Pa 231.036	92 U 238.029	93 Np 237.048	94 Pu (244)	95 Am (243)	96 Cm (247)	97 Bk (247)	98 Cf (251)	99 Es (252)	100 Fm (257)	101 Md (258)	102 No (259)	103 Lr (260)

\* At the time of writing, element 112 had been given the suggested name Copernicium.

FIGURE 2.4 The periodic table of the elements.

chemical properties. This is especially true in groups IIIA, IVA, VA, VIA, and VIIA. For example, nitrogen bears little chemical resemblance to bismuth.

## 2.2.2 Ionization Energy

An important property of atoms that is related to their chemical behavior is the *ionization potential* or *ionization energy*. In general, ionization energy can be defined as the energy needed to remove an electron from a gaseous atom. For hydrogen, there is only one ionization potential because the atom has only one electron. Atoms having more than one electron have an ionization potential for each electron, and these often differ markedly. After the first electron is removed, succeeding electrons are removed from an ion that is already positively charged. The series of ionization energies ( $I$ ) for a given atom increases as  $I_1 < I_2 < \dots < I_n$ .

Ionization energies can be measured directly to provide evidence for the ordering of the energy levels in atoms. Figure 2.5 shows the variation in first ionization energy with position of atoms in the periodic table.

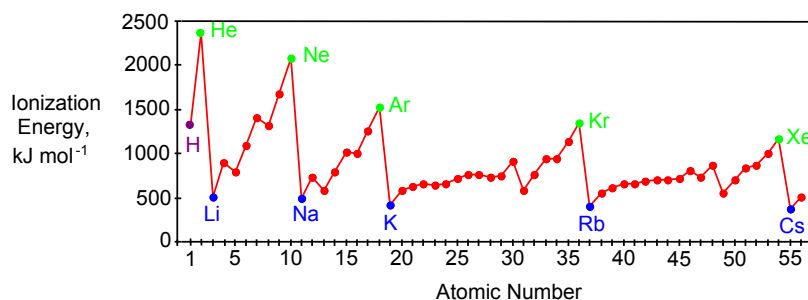


FIGURE 2.5 Ionization energy as a function of atomic number.

An extensive table of ionization energies is given in Appendix B. Although the energy necessary to remove several electrons from multielectron atoms can be determined, usually no more than three or four are removed when compounds form. As a result, oxidation states as high as seven (e.g., Mn in  $\text{MnO}_4^-$ ) are common, but such species do not contain atoms that have lost seven electrons. Consequently, the table presented in Appendix B shows only the first three ionization energies for atoms up to atomic number 55, and only the first two are given for heavier atoms.

The graph of the ionization energies as a function of atomic number shown in [Figure 2.5](#) reveals a number of useful generalizations that will now be described.

1. The highest first ionization energy, about  $2400 \text{ kJ mol}^{-1}$ , is for He. As a group, the noble gases have the highest ionization energies and the alkali metals have the lowest.
2. The first ionization energy shows a decrease as one goes down a given group. For example, Li, 513.3; Na, 495.8; K, 418.8; Rb, 403; Cs, 375.7  $\text{kJ mol}^{-1}$ . This trend is to be expected because even though nuclear charge increases, so does the extent of shielding by inner shell electrons. Electrons in the inner shells effectively screen outer electrons from part of the attraction to the nucleus. Going down the group of elements, the outside electrons lost in ionization are farther away from the nucleus, and the other groups show a similar trend.
3. For some elements, the first ionization energy alone is not always relevant because the elements may not exhibit a stable oxidation state of +1. For example, in Group IIA, the sum of the first two ionization energies should be compared because the +2 ions are more common. The values are as follows: Be, 2656.5; Mg, 2188.4; Ca, 1734.7; Sr, 1613.7; and Ba, 1467.9  $\text{kJ mol}^{-1}$ .
4. The effect of closed shells is apparent. For example, sodium has a first ionization energy of only  $495.8 \text{ kJ mol}^{-1}$  whereas the second is  $4562.4 \text{ kJ mol}^{-1}$ . The second electron removed comes from  $\text{Na}^+$ , and it is removed from the filled  $2p$  shell. For Mg, the first two ionization potentials are 737.7 and  $1450.7 \text{ kJ mol}^{-1}$ , and the difference represents the additional energy necessary to remove an electron from a +1 ion. Thus, the enormously high second ionization energy for Na is largely due to the closed shell effect.
5. There is a general increase in first ionization energy as one goes to the right across a row in the periodic table. This increase is a result of the increase in nuclear charge and a general size decrease.
6. The first ionization energy for N is slightly higher than that for O. This is a manifestation of the effect of the stability of the half-filled shell in N. As a result of the oxygen atom having one electron beyond a half-filled shell, the first electron of oxygen is easier to remove. A similar effect is seen for P and S, although the difference is smaller than it is for N and O. As one goes farther down in the periodic table, the effect becomes less until it disappears when the ionization energy for Sb and Te are compared.

### 2.2.3 Electron Affinity

Many atoms have a tendency to add one or more electrons when forming compounds. In most cases, this is an energetically favorable process. As will be described in Chapter 4, one step in the formation of an ionic bond is the addition of an electron to a neutral, gaseous atom to give a negative ion, which can be shown as



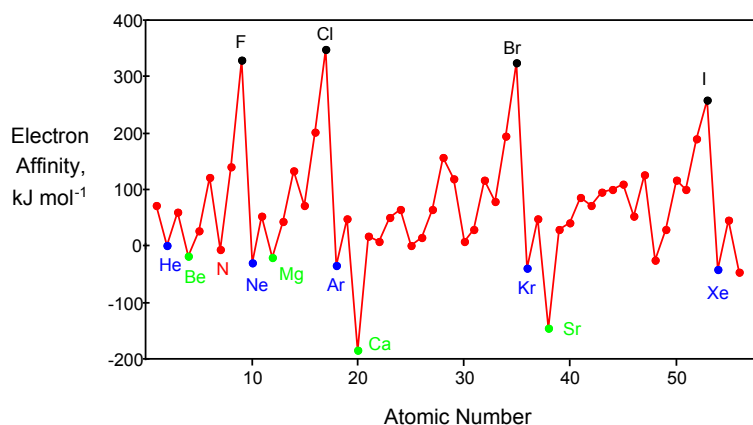
The addition of an electron to an uncharged atom or negatively charged ion is referred to as the *electron addition enthalpy*. The energy associated with *removal* of an electron from a negatively charged species (the atom that has gained an electron) is the *electron affinity*.



In most cases, the *enthalpy* associated with this process is *positive* meaning that energy is required to remove the electron from the atom that has gained it. Most atoms add one electron with the release of energy, but when  $\text{O}^{2-}$  and  $\text{S}^{2-}$  are formed, the atom must add *two* electrons. The addition of a second electron is always unfavorable. There is no atom that will add two electrons with a release of energy. Therefore, in forming compounds that contain such ions there must be some other factor that makes the process energetically favorable.

Experimentally, the electron affinity is difficult to measure, and most of the tabulated values are obtained from thermochemical cycles where the other quantities are known (see Chapter 4). Electron affinities are often given in units other than those needed for a particular use. Therefore, it is useful to know that  $1 \text{ eV molecule}^{-1} = 23.06 \text{ kcal mol}^{-1}$ , and  $1 \text{ kcal} = 4.184 \text{ kJ}$ . Electron affinities for many nonmetallic atoms are shown in [Table 2.3](#).

Process	Electron Affinity, $\text{kJ mol}^{-1}$
$\text{H}^-(g) \rightarrow \text{H}(g) + e^-(g)$	72.8
$\text{F}^-(g) \rightarrow \text{F}(g) + e^-(g)$	328
$\text{Cl}^-(g) \rightarrow \text{Cl}(g) + e^-(g)$	349
$\text{Br}^-(g) \rightarrow \text{Br}(g) + e^-(g)$	324.7
$\text{I}^-(g) \rightarrow \text{I}(g) + e^-(g)$	295.2
$\text{B}^-(g) \rightarrow \text{B}(g) + e^-(g)$	26.7
$\text{C}^-(g) \rightarrow \text{C}(g) + e^-(g)$	121.9
$\text{N}^-(g) \rightarrow \text{N}(g) + e^-(g)$	-7
$\text{O}^-(g) \rightarrow \text{O}(g) + e^-(g)$	141
$\text{S}^-(g) \rightarrow \text{S}(g) + e^-(g)$	200.4
$\text{O}^{2-}(g) \rightarrow \text{O}^-(g) + e^-(g)$	-845
$\text{S}^{2-}(g) \rightarrow \text{S}^-(g) + e^-(g)$	-531



**FIGURE 2.6** Electron affinity as a function of atomic number.

There are several interesting comparisons of electron affinities. The first is that F has a *lower* electron affinity than Cl. The fact that F is such a small atom and the added electron must be in close proximity to the other seven valence shell electrons is the reason. Below Cl in the periodic table, there is a decrease in electron affinity as one goes down in the remainder of the group:  $\text{Cl} > \text{Br} > \text{I}$ , in accord with the increase in size. In a general way, there is an increase in electron affinity as one goes to the right in a given row in the periodic table. This is the result of the increase in nuclear charge, but the electron affinity of nitrogen ( $-7 \text{ kJ mol}^{-1}$ ) appears to be out of order in the first long row. This is a result of the stability of the half-filled shell, and the oxygen atom having one electron beyond a half-filled  $2p$  shell. Group IIA elements ( $ns^2$ ) and the noble gases ( $ns^2 np^6$ ) have negative values as a result of the filled shell configurations. Figure 2.6 shows the trend in electron affinity graphically as a function of atomic number. Note that the highest values correspond to the Group VIIA elements.

## 2.2.4 Electronegativity

When two atoms form a covalent bond, they do not share the electrons equally unless the atoms are identical. The concept of *electronegativity* was introduced by Linus Pauling to explain the tendency of an atom in a molecule to attract electrons. The basis for Pauling's numerical scale that describes this property lies in the fact that polar covalent bonds between atoms

of different electronegativity are more stable than if they were purely covalent. The stabilization of the bond,  $\Delta_{AB}$ , in a diatomic molecule AB due to this effect can be expressed as

$$\Delta_{AB} = D_{AB} - (1/2)[D_{AA} + D_{BB}] \quad (2.15)$$

where  $D_{AA}$  and  $D_{BB}$  represent the bond energies in the diatomic species  $A_2$  and  $B_2$ , respectively, and  $D_{AB}$  is the bond energy of the molecule AB. Thus, the term  $\Delta_{AB}$  represents the additional contribution to the A–B bond strength as a result of the atoms having different electronegativities. The extent of the stabilization can also be expressed in terms of the difference in the electronegativities of the atoms by the equation

$$\Delta_{AB}(\text{kJ mol}^{-1}) = 96.48|\chi_A - \chi_B|^2 \quad (2.16)$$

where  $\chi_A$  and  $\chi_B$  are the electronegativities for atoms A and B. Therefore, it is the *difference* between the electronegativities that is related to the additional stabilization of the bond, but some value for the electronegativity for at least one atom had to be specified. Assigning a value for one atom leads to a *relative* value for each other atom. The Pauling electronegativity scale was established with fluorine being given a value of 4.0, and the other atoms then have values between 0 and 4. Table 2.4 shows electronegativity values for several atoms.

The electronegativity scale established by Pauling is not the only such scale, and the electronegativity of an atom A has been defined by Mulliken as

$$\chi_A = (1/2)[I + E] \quad (2.17)$$

where  $I$  and  $E$  are the ionization potential and electron affinity of the atom. This is a reasonable approach because the ability of an atom in a molecule to attract electrons would be expected to be related to the ionization potential and electron affinity. Both of these properties are also related to the ability of an atom to attract electrons. Most electronegativities on the Mulliken scale differ only slightly from the Pauling values. For example, fluorine has the Pauling electronegativity of 4.0 and a value of 3.91 on the Mulliken scale. A different approach was used by Allred and Rochow to establish an electronegativity scale. This scale is based on a consideration of the electrostatic force holding a valence shell electron in an atom of radius,  $r$ , by an effective nuclear charge,  $Z^*$ . This electronegativity value,  $\chi_{AR}$ , is given by

$$\chi_{AR} = 0.359(Z^*/r^2) + 0.744 \quad (2.18)$$

Many other electronegativity scales have been developed, but the three scales described are ones most frequently used, and qualitative agreement between the scales is quite good. One of the most important uses of electronegativity values is in deciding bond polarities and in estimating the importance of possible resonance structures for molecules. For example, based on electronegativities, HCl should have hydrogen at the *positive* end of the dipole and chlorine at the *negative* end. In drawing structures for molecules, it will be observed that those structures corresponding to an accumulation of electron density on atoms of high electronegativity are usually more important. This situation will be treated more fully in the next chapter.

**TABLE 2.4** Electronegativities of Atoms

H 2.2									
Li 1.0	Be 1.6				B 2.0	C 2.6	N 3.0	O 3.4	F 4.0
Na 1.0	Mg 1.3				Al 1.6	Si 1.9	P 2.2	S 2.6	Cl 3.2
K 0.8	Ca 1.0	Sc 1.2	...	Zn 1.7	Ga 1.8	Ge 2.0	As 2.2	Se 2.6	Br 3.0
Rb 0.8	Sr 0.9	Y 1.1	...	Cd 1.5	In 1.8	Sn 2.0	Sb 2.1	Te 2.1	I 2.7
Cs 0.8	Ba 0.9	La 1.1	...	Hg 1.5	Tl 1.4	Pb 1.6	Bi 1.7	Po 1.8	At 2.0

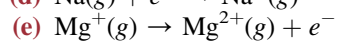
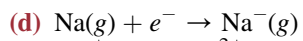
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**PROBLEMS**

- Write the set of four quantum numbers for the “last” electron in each of the following.
  - Li
  - Ca
  - Sc
  - Fe
- Write all the possible sets of four quantum numbers for electrons in the  $5d$  subshell.
- Write complete electron configurations for the following atoms.
  - O
  - Kr
  - Ni
  - Ti
  - Fr
- Write complete electron configurations for the following ions.
  - $\text{Co}^{3+}$
  - $\text{Sn}^{4+}$
  - $\text{N}^{3-}$
  - $\text{Se}^{2+}$
  - $\text{Fe}^{3+}$
- Write complete electron configurations for the following ions.
  - $\text{Mo}^{2+}$
  - $\text{Cu}^+$
  - $\text{S}^{2-}$
  - $\text{Mg}^{2+}$
  - $\text{I}^-$
- Explain why atoms such as Cr and Cu do not have “regular” electron configurations.
- For each of the following pairs, predict which species would have the higher first ionization potential.
  - Na or Al
  - Ca or Ba
  - Br or Kr
  - Fe or Cl
  - C or N
- Explain why the first ionization potential for Be is slightly higher than that of B.
- Explain why the noble gases have the highest first ionization potentials.
- For each of the following tell whether the process would be exothermic or endothermic and provide a brief explanation.
  - $\text{K}(g) \rightarrow \text{K}^+(g) + e^-$
  - $\text{Cl}(g) + e^- \rightarrow \text{Cl}^-(g)$
  - $\text{O}^-(g) + e^- \rightarrow \text{O}^{2-}(g)$



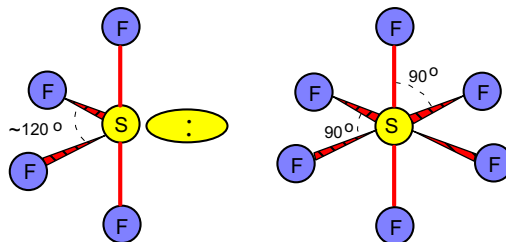


11. For each of the following pairs, predict which species would have the higher electron affinity.
- (a) Cl or I
  - (b) F or Ne
  - (c) B or C
  - (d) O or  $\text{O}^-$
12. Explain why the electron affinity for nitrogen is much lower than that of either carbon or oxygen.

## Chapter 3

# Covalent Bonding and Molecular Structure

Because so much of the chemical behavior of molecules is related to their structures, the study of descriptive chemistry must include a discussion of molecular structure. The reasons for this are quite simple and straightforward. For example, many of the chemical characteristics of nitrogen are attributable to the stable structure of the  $N_2$  molecule,  $:N\equiv N:$ . The triple bond in the  $N_2$  molecule is very strong and disruption of the bond requires considerable energy. Thus, the bond strength of the  $N_2$  molecule is responsible for many chemical properties of nitrogen (such as it being a relatively unreactive gas). Likewise, to understand the basis for the enormous difference in the chemical behavior of  $SF_4$  and  $SF_6$ , it is necessary to understand the difference between the structures of these molecules, which can be shown as



Moreover, to understand why  $SF_6$  exists as a stable compound whereas  $SCl_6$  does not, we need to know something about the properties of the S, F, and Cl atoms. As another illustration, it may be asked why the  $PO_4^{3-}$  ion is quite stable but  $NO_4^{3-}$  is not. Throughout this *descriptive* chemistry book, reference will be made in many instances to differences in chemical behavior that are based on atomic and molecular properties. Certainly not all chemical characteristics are predictable from an understanding of atomic and molecular structure. However, structural principles are useful in so many cases (for both comprehension of facts and prediction of properties) that a study of molecular structure is essential.

### 3.1 MOLECULAR STRUCTURE

There are two principal approaches to describe bonding in molecules by quantum mechanical methods. These are known as the *valence bond* method and the *molecular orbital* (MO) method. Basically, the difference is in the way in which molecular wave functions are expressed. The valence bond method has as an essential feature that atoms retain their individuality, and the molecule arises from bringing together complete atoms. In the MO method, the nuclei are brought to their positions in the molecule and the electrons are placed in MOs that encompass the whole molecule. The valence bond method is older and follows quite naturally the notion of two atoms combining to form a molecule by sharing of electrons in atomic orbitals. In this section, bonding in diatomic molecules will be described.

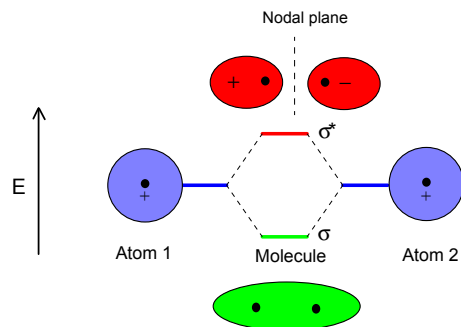
#### 3.1.1 Molecular Orbitals

In the MO approach, atomic orbitals lose their identities as they form orbitals encompassing the whole molecule. Wave functions for MOs can be constructed from atomic wave functions by taking linear combinations. If the atomic wave functions are represented by  $\phi_1$  and  $\phi_2$ , the molecular wave functions,  $\psi_b$  and  $\psi_a$ , can be written as the combinations

$$\psi_b = a_1\phi_1 + a_2\phi_2 \quad (3.1)$$

and

$$\psi_a = a_1\phi_1 - a_2\phi_2 \quad (3.2)$$



**FIGURE 3.1** Molecular orbitals and electron density contours formed from the combination of two  $s$  atomic orbitals.

where  $a_1$  and  $a_2$  are constants. The *square* of the wave function is related to probability of finding electrons. Squaring both sides of the equations shown above gives

$$\psi_b^2 = a_1^2\phi_1^2 + a_2^2\phi_2^2 + 2a_1a_2\phi_1\phi_2 \quad (3.3)$$

$$\psi_a^2 = a_1^2\phi_1^2 + a_2^2\phi_2^2 - 2a_1a_2\phi_1\phi_2 \quad (3.4)$$

The term  $a_1^2\phi_1^2$  represents the probability of finding electrons from atom 1 and  $a_2^2\phi_2^2$  is the probability from atom 2. A covalent bond can be defined as the increased probability of finding electrons between two atoms resulting from electron sharing. As shown in Eq. (3.3), the term  $2a_1a_2\phi_1\phi_2$  is proportional to the *increased* probability of finding electrons between the atoms caused by the bond between them. In Eq. (3.4), the term  $-2a_1a_2\phi_1\phi_2$  leads to a *decreased* probability of finding electrons between the two atoms. In fact, there is a nodal plane between them where the probability goes to zero.

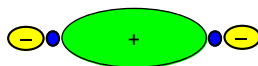
The energy state corresponding to  $\psi_b$  is called the *bonding* state and that arising from  $\psi_a$  is the *antibonding* state. Two MOs have resulted from the combination of two atomic orbitals. The energy level diagram for the atomic and molecular states is shown in Figure 3.1.

For  $H_2$ , the two electrons can be placed in the bonding state to give the configuration  $\sigma^2$ . For a covalent bond, the *bond order*,  $B$ , is defined as

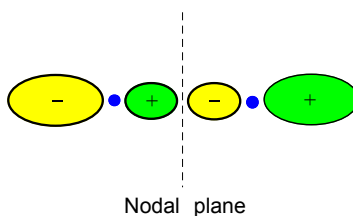
$$B = \frac{N_b - N_a}{2} \quad (3.5)$$

where  $N_b$  and  $N_a$  are the numbers of electrons in bonding and antibonding orbitals, respectively. For the  $H_2$  molecule,  $B = 1$ , which describes a single bond. Combination of two wave functions for  $2s$  orbitals gives a similar result.

When the combinations of  $2p$  orbitals are considered, there are two possible results. The bond is presumed to lie along the  $z$ -axis so the  $p_z$  orbitals may combine “end on” to give either a  $\sigma$  bond,



or the combination of atomic wave functions can also give rise to the  $\sigma^*$  antibonding state that is represented as follows.



If the  $p_z$  orbitals combine end on, the remaining  $p$  orbitals ( $p_x$  and  $p_y$ ) must combine side-to-side so that two  $\pi$  bonds are formed. Two antibonding states also result, and the  $\pi$  and  $\pi^*$  orbitals are shown in Figure 3.2. Note that four atomic wave functions have resulted in four molecular wave functions, two that are bonding and two that are antibonding in character.

The energy level diagram that results from combinations of  $2s$  and  $2p$  orbitals from the two atoms is shown in Figure 3.3. It should be mentioned that in representing the orbitals as shown in Figure 3.3(a), it is assumed that the  $2s$  orbital lies lower in energy than the  $2p_z$  orbital and that the difference is sufficiently large that there is no interaction between them. If the energy difference between them is relatively small, the  $2s$  and  $2p_z$  orbital can interact (through partial

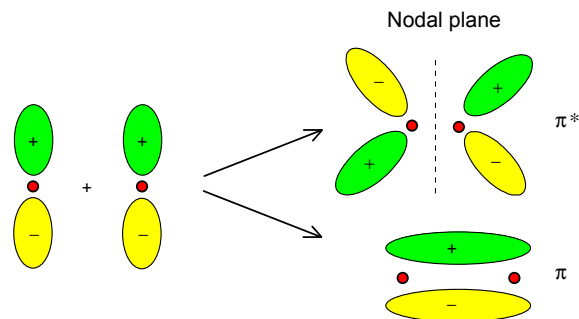


FIGURE 3.2 Electron density contours of bonding and antibonding orbitals formed from two atomic  $p$  orbitals.

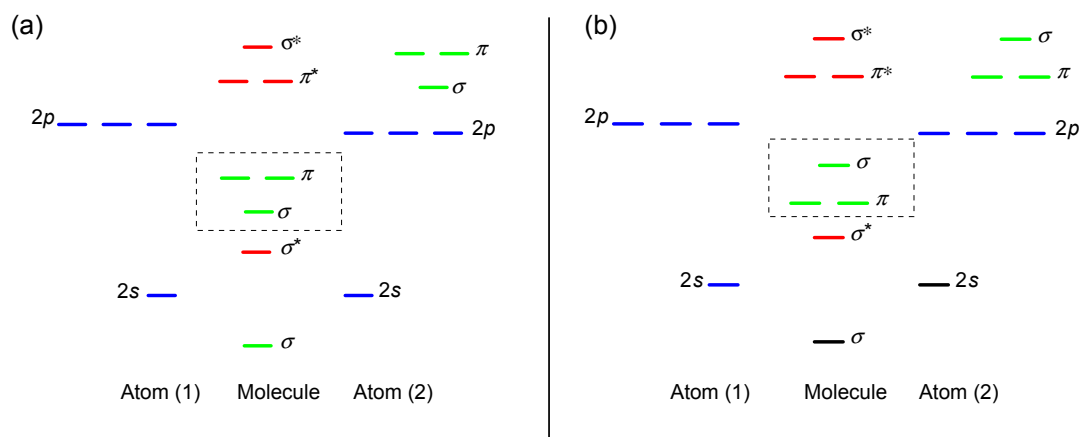


FIGURE 3.3 Molecular orbital diagrams that result from combining  $2s$  and  $2p$  atomic orbitals. In (b), there is hybridization of the  $2s$  and  $2p_z$  orbitals that causes the  $\pi$  orbitals arising from the  $2p_x$  and  $2p_y$  orbitals to lie lower in energy than the  $\sigma_{2p}$ . In (a), the energy difference is presumed to be large enough to prevent hybridization and the  $\sigma_{2s}$  lies lower in energy than the  $\pi$  orbitals.

hybridization), which results in a change in the energies of the resulting MOs. This result causes the  $\sigma_{2p}$  MO to lie higher in energy than the  $\pi$  orbitals that arise from combining the  $p_x$  and  $p_y$  orbitals. In this case, the orbital energy diagram is as shown in Figure 3.3(b).

The difference in energy between the  $2s$  and  $2p_z$  orbitals changes markedly in the second period of the periodic table. For example, the difference in energy between the  $2s$  and  $2p_z$  orbitals in Li is about 1.85 eV (178 kJ mol<sup>-1</sup>) but in F, the difference is about 20 eV (1930 kJ mol<sup>-1</sup>). Therefore, hybridization of the  $2s$  and  $2p_z$  orbitals occurs early in the period which causes the arrangement of orbitals shown in Figure 3.3(b) to be correct for diatomic molecules of B<sub>2</sub>, C<sub>2</sub>, and N<sub>2</sub>.

For the B<sub>2</sub> molecule, which has six valence shell electrons, populating the orbitals as shown in Figure 3.3(a) would lead to the configuration  $\sigma_{2s}^2 \sigma_{2s}^{*2} \sigma_{2p_z}^2$  and the molecule would be diamagnetic. On the other hand, if the orbitals are arranged as shown in Figure 3.3(b), the configuration would be  $\sigma_{2s}^2 \sigma_{2s}^{*2} \pi_{2p_y}^1 \pi_{2p_z}^1$  and the molecule would be paramagnetic. Because the B<sub>2</sub> molecule is paramagnetic, Figure 3.3(b) is the correct energy level diagram for B<sub>2</sub>. The C<sub>2</sub> molecule has eight valence shell electrons. Populating the MO as shown in Figure 3.3(a) would yield  $\sigma_{2s}^2 \sigma_{2s}^{*2} \sigma_{2p_z}^2 \pi_{2p_x}^1 \pi_{2p_y}^1$ , indicating that the molecule would be paramagnetic. However, the C<sub>2</sub> molecule is diamagnetic, which is consistent with the orbital diagram shown in Figure 3.3(b). Later in the period, hybridization does not occur and Figure 3.3(a) gives the correct order of MOs for O<sub>2</sub> and F<sub>2</sub>.

One of the interesting successes of the MO approach to bonding in diatomic molecules is the fact that molecules such as O<sub>2</sub> and B<sub>2</sub> are correctly predicted to be paramagnetic, but the valence bond structures for these molecules are unsatisfactory. Properties for many diatomic species are shown in Table 3.1.

A consideration of the data shown in Table 3.1 reveals some interesting and useful relationships. In general, as the bond order increases, the bond energy increases, and the bond length,  $r$ , decreases. These trends will be observed in later chapters.

**TABLE 3.1** Properties for Diatomic Molecules

Molecule	$N_b$	$N_a$	$B^a$	$r$ , pm	Bond Energy, $\text{kJ mol}^{-1}$
$\text{H}_2^+$	1	0	0.5	106	255.7
$\text{H}_2$	2	0	1	74	458.3
$\text{He}_2^+$	2	1	0.5	108	299.0
$\text{Li}_2$	2	0	1	262	99.4
$\text{B}_2$	4	2	1	159	289.4
$\text{C}_2$	6	2	2	131	569.2
$\text{N}_2$	8	2	3	109	941.6
$\text{O}_2$	8	4	2	121	490.1
$\text{O}_2^+$	8	3	2.5	112	623.3
$\text{O}_2^-$	8	5	1.5	133	~275
$\text{O}_2^{2-}$	8	6	1	149	142
$\text{F}_2$	8	6	1	142	154.4
$\text{N}_2^+$	7	2	2.5	112	836.5
BN	6	2	2	128	385.8
BO	7	2	2.5	120	771.9
CN	7	2	2.5	118	786.2
CO	8	2	3	113	1071.1
NO	8	3	2.5	115	677.4
$\text{NO}^+$	8	2	3	106	1020
SO	8	4	2	149	497.9
PN	8	2	3	149	577.0
SiO	8	2	3	151	773.6
LiH	2	0	1	160	241.4
NaH	2	0	1	189	192.9
PO	8	3	2.5	145	510.5

One eV/molecule = 23.06 kcal mol<sup>-1</sup> = 96.48 kJ mol<sup>-1</sup>.  
<sup>a</sup>B is the bond order.

### 3.1.2 Orbital Overlap

A covalent bond arises from the sharing of electrons between atoms. This results in an increase in electron density between the two atoms. Thus, covalent bonds are represented as the *overlap of atomic orbitals*. The overlap between two atomic orbitals,  $\phi_1$  and  $\phi_2$ , on atoms 1 and 2 is represented in terms of the *overlap integral*,  $S$ , which is defined as

$$S = \int \phi_1 \phi_2 d\tau. \quad (3.6)$$

When an infinite distance separates the two atoms, the overlap of the orbitals is zero. If the orbitals overlap completely, as when the atoms are superimposed, then  $S = 1$ . Thus, the overlap integral varies between 0 and 1, and the greater the value, the greater the extent of orbital overlap.

As two atoms are brought together, the orbitals may interact in different ways. First, if the orbitals have the correct mathematical symmetry, there is a reinforcement or enhanced probability of finding the electrons between the atoms in the regions where the orbitals overlap as shown in [Figure 3.4](#), and  $S > 0$ . These cases are referred to as *bonding overlap*.

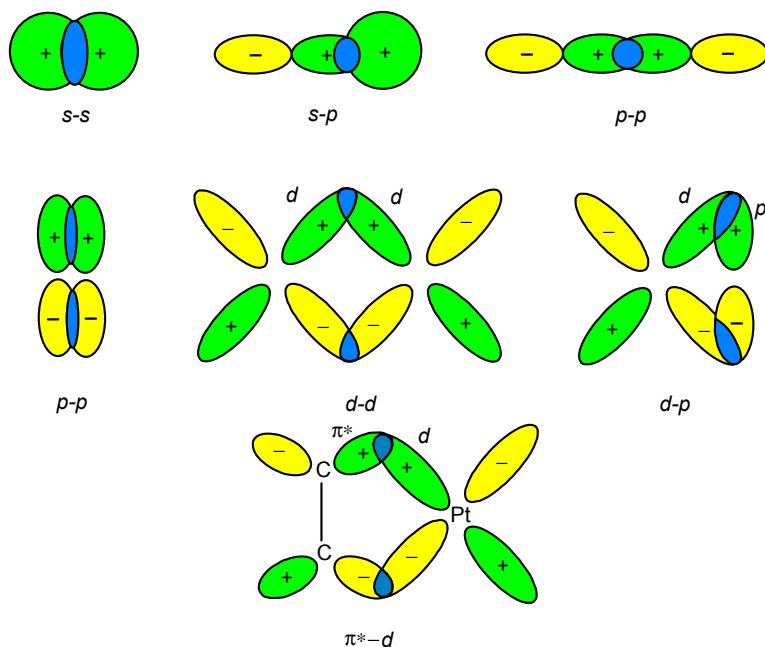


FIGURE 3.4 Some types of overlap of orbitals that can lead to bond formation ( $S > 0$ ).

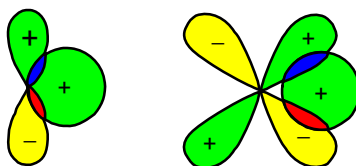


FIGURE 3.5 Nonbonding arrangements of orbitals ( $S = 0$ ). The regions of overlap involve orbitals of opposite mathematical sign so each favorable overlap is cancelled, so there is no *net* overlap.

Second, in some cases, as shown in Figure 3.5, the orbitals overlap so that there is favorable overlap in one region that is cancelled in another. The result is that  $S = 0$  and there is no overall increased probability of finding electrons shared between the two atoms. That is, the wave functions are said to be *orthogonal*, and these cases are referred to as *nonbonding*.

The last type of orbital interaction is shown in Figure 3.6. In these cases, the orbitals or their lobes have opposite signs so that there is a *decreased* probability of finding the electrons between the two atoms. These situations are referred to as *antibonding* cases.

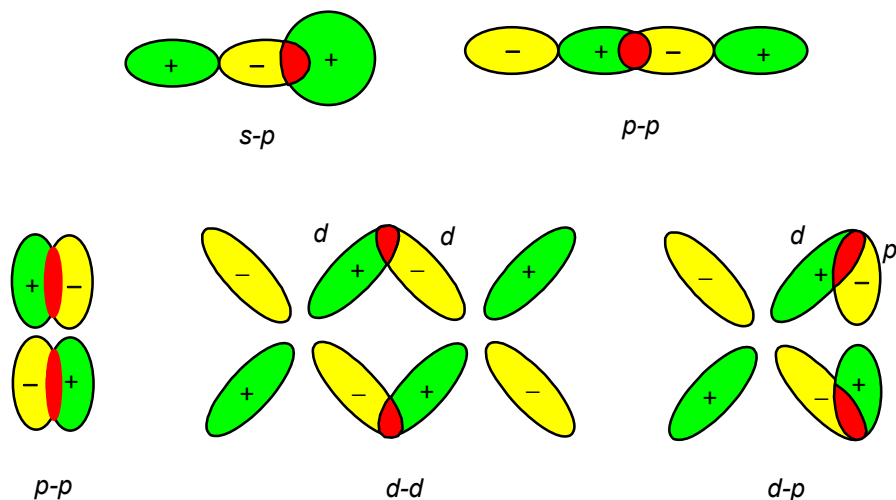


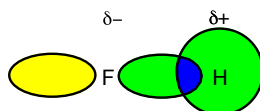
FIGURE 3.6 Some antibonding arrangements of orbitals.

Because the degree to which electron density increases between two nuclei depends on the value of the overlap integral, covalent bond strength essentially depends on that quantity. Thus, conditions that lead to larger overlap integral values will generally increase the bond strength. A large value of the overlap integral occurs under the following conditions:

1. Orbitals should have similar sizes (and hence energies) for effective overlap.
2. The atoms should be positioned in the molecule in such a way that effective overlap can occur with a minimum of orbital distortion.
3. Effective overlap (overlap integral  $> 0$ ) occurs between orbitals that give the correct symmetry combinations (positive with positive and negative with negative regions).

### 3.1.3 Polar Molecules

The shared electron pair in HF is distorted toward the F end of the molecule because of the electronegativity difference between H and F atoms. As a result of the electrons not being shared equally, the structure can be shown as



where  $\delta+$  and  $\delta-$  are some small increment of positive and negative charge, respectively. If the electrons were shared equally, the result would be H:F and there would be no charge separation in the molecule. If the electron were *completely* transferred from the H atom to the F atom, the structure could be shown as  $H^+ F^-$  which represents an ionic bond.

The *dipole moment* is a way of expressing nonsymmetrical charge distribution of electrons in a molecule. It is represented as  $\mu$ , which is defined by the equation

$$\mu = q \times r \quad (3.7)$$

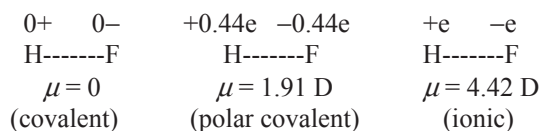
where  $q$  = the charge separated in esu ( $1 \text{ esu} = 1 \text{ g}^{1/2} \text{ cm}^{3/2} \text{ sec}^{-1}$ ) and  $r$  = the distance of separation in cm. For H–F, the actual bond length is 92 pm ( $0.92 \text{ \AA}$ ). If the structure was totally ionic, the amount of charge separated would represent the charge of an electron,  $4.8 \times 10^{-10}$  esu. Thus, the dipole moment,  $\mu_1$  would be

$$\mu_1 = 4.8 \times 10^{-10} \text{ esu} \times 0.92 \times 10^{-8} \text{ cm} = 4.42 \times 10^{-18} \text{ esu cm} = 4.42 \text{ D}$$

where 1 Debye = 1 D =  $10^{-18}$  esu cm. The actual dipole moment for HF is 1.91 D so only a fraction of the electron charge is transferred from H to F. The actual quantity of charge separated can be calculated from the relationship shown in Eq. (3.7).

$$1.91 \times 10^{-18} \text{ esu cm} = q \times 0.92 \times 10^{-8} \text{ cm}$$

Therefore, the actual value of  $q$  is  $2.1 \times 10^{-10}$  esu. The fraction of an electron that appears to have been transferred is  $2.1 \times 10^{-10} \text{ esu} / 4.8 \times 10^{-10} \text{ esu} = 0.44$  of the electron charge so it appears that 44% of the electron has been transferred. The following structures illustrate this situation.



Only the middle structure exists, and it is sometimes said that the bond in HF is 44% ionic. What is really meant is that the true structure behaves as though it is a composite of 56% of the nonpolar structure and 44% of the ionic structure. The *true* structure is a *resonance hybrid* of these hypothetical structures, neither of which actually exists.

The contributions of covalent and ionic structures to a molecular wave function for a bond in a diatomic molecule can be expressed in terms of

$$\Psi_{\text{molecule}} = \Psi_{\text{covalent}} + \lambda \Psi_{\text{ionic}} \quad (3.8)$$

where  $\lambda$  is a constant known as weighting factor that needs to be determined. In the case of HF, a purely covalent structure with equal sharing of the bonding pair of electrons would result in a zero dipole moment for the molecule as shown above. A strictly ionic structure would produce a dipole moment of 4.42 D. The ratio of the observed dipole moment,  $\mu_{\text{obs}}$ , to that

calculated for an ionic structure,  $\mu_{\text{ionic}}$ , gives the fraction of ionic character of the bond. Therefore, the percent ionic character is given as

$$\% \text{ Ionic character} = 100 \left( \frac{\mu_{\text{obs}}}{\mu_{\text{ionic}}} \right) \quad (3.9)$$

It should be recalled that it is the square of the coefficients that is related to the weighting given to the structure (see Eq. (3.3)). Consequently,  $\lambda^2$  is related to the weighting given to the ionic structure, and  $1^2 + \lambda^2$  is the total contribution of both the covalent and ionic structures. Therefore, the ratio of the weighting of the ionic structure to the total is  $\lambda^2/(1^2 + \lambda^2)$  and

$$\% \text{ Ionic character} = \frac{100 \lambda^2}{(1^2 + \lambda^2)} \quad (3.10)$$

so that

$$\frac{\lambda^2}{1^2 + \lambda^2} = \frac{\mu_{\text{obs}}}{\mu_{\text{ionic}}} \quad (3.11)$$

For HF,  $\mu_{\text{obs}}/\mu_{\text{ionic}} = 0.44$  from which  $\lambda$  can be calculated to be 0.87.

At this point, two additional equations should be presented because they are sometimes used to estimate the relative contributions of covalent and ionic structures. These semiempirical relationships are based on electronegativity,  $\chi$ , and are as follows.

$$\% \text{ Ionic character} = 16|\chi_{\text{A}} - \chi_{\text{B}}| + 3.5|\chi_{\text{A}} - \chi_{\text{B}}|^2 \quad (3.12)$$

$$\% \text{ Ionic character} = 18|\chi_{\text{A}} - \chi_{\text{B}}|^{1.4} \quad (3.13)$$

Although these equations appear quite different, the predicted percentage of ionic character is approximately the same for  $|\chi_{\text{A}} - \chi_{\text{B}}|$  in the range of 1–2. Using these equations and Eq. (3.10), it is possible to estimate the value of  $\lambda$  in Eq. (3.10) if the electronegativities of the atoms are known.

Although covalent bonds in which there is unequal sharing of electrons can be treated as vector quantities, the combining of those vectors to calculate an overall dipole moment for the molecule is not always successful. Unshared pairs of electrons also affect the molecular polarity to a considerable degree. However, in order to interpret the physical properties and many aspects of the chemical behavior of molecules, it is necessary to understand their polarity. Consequently, Table 3.2 shows the dipole moments of many common inorganic molecules.

It is a fundamental principle that when MOs are formed from atomic orbitals that have different energies, the bonding orbital retains more of the character of the atomic orbital having lower energy. In other words, there is not an exactly equal “mix” of the orbitals except when the atomic orbitals have the same energy. The greater the electronegativity difference, the more ionic the bond becomes, and the MO in that case represents an orbital on the atom having *higher* electronegativity (which gains the electron). Figure 3.7 shows these observations in terms of the MO energies. The first case represents a purely covalent bond, the second a polar covalent bond, and the third case a bond that is essentially ionic.

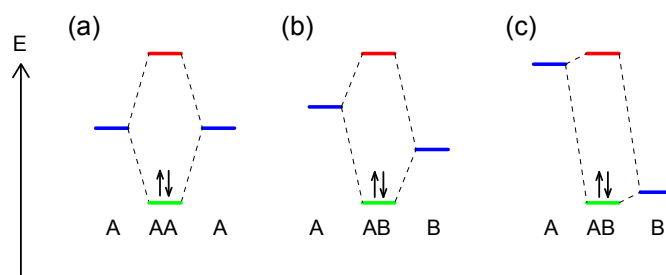
In Figure 3.7(a), the atoms have the same electronegativity (as in  $\text{H}_2$ ). In (b), atom B is more electronegative and the bonding orbital lies closer to the atomic energy level of atom B (as in HF or HCl). In (c), the difference in electronegativity between atoms A and B is large enough that the MO is essentially an atomic orbital on B. The effect is that the electron is transferred to the atom of higher electronegativity (as in LiF).

### 3.1.4 Geometry of Molecules Having Single Bonds

Although the MO description of bonding has some mathematical advantages, simple valence bond representations of structures are adequate for many purposes. The structures of molecules that have only single bonds (and in some cases unshared pairs of electrons on the central atom) are based on placing the electrons in orbitals that minimize repulsion. When drawing the structures of molecules, a solid line or dash represents a *pair* of electrons (either shared or unshared). In other cases, electrons are shown as individual dots and because teachers use both schemes, we do also in this chapter. In some cases to keep from cluttering the drawing, unshared pairs of electrons are not shown on peripheral atoms.

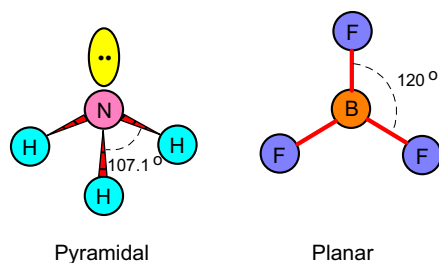


Molecule	Dipole Moment, D	Molecule	Dipole Moment, D
H <sub>2</sub> O	1.85	NH <sub>3</sub>	1.47
H <sub>2</sub> S	0.97	HN <sub>3</sub>	1.70
PH <sub>3</sub>	0.58	AsH <sub>3</sub>	0.20
SbH <sub>3</sub>	0.12	AsCl <sub>3</sub>	1.59
AsF <sub>3</sub>	2.59	HF	1.82
NF <sub>3</sub>	0.23	HCl	1.08
PCl <sub>3</sub>	0.78	HBr	1.43
ClF <sub>3</sub>	0.60	HI	0.44
SOCl <sub>2</sub>	1.45	F <sub>2</sub> NH	1.92
SO <sub>2</sub> Cl <sub>2</sub>	1.81	SO <sub>2</sub>	1.63
OPF <sub>3</sub>	1.76	SPF <sub>3</sub>	0.64
SF <sub>4</sub>	0.63	IF <sub>5</sub>	2.18
HNO <sub>3</sub>	2.17	H <sub>2</sub> O <sub>2</sub>	2.2
N <sub>2</sub> H <sub>4</sub>	1.75	F <sub>2</sub> O	0.30
NO	0.16	NO <sub>2</sub>	0.32
N <sub>2</sub> O	0.16	PFC <sub>4</sub>	0.21
OCS	0.72	SeO <sub>2</sub>	2.62



**FIGURE 3.7** Molecular orbitals showing the effects of electronegativity. (a) The two atoms have the same electronegativity. (b) Atom B has higher electronegativity. The MO has more of the character of an orbital of atom B. (c) The difference in electronegativity is large enough that the electron pair resides in an orbital that is essentially one of atom B (ionic bond).

It is first necessary to determine the number of electrons surrounding the central atom. That number includes the number of valence shell electrons the atom has (indicated by the group number) and those contributed by the peripheral atoms. For example, in NH<sub>3</sub> there are eight electrons around the nitrogen atom (five from N and one each from three H atoms). Those electrons occupy four orbitals that point toward the corners of a tetrahedron. In BF<sub>3</sub>, there are six electrons around the boron atom, three from that atom and one from each F atom. To minimize repulsion, the three pairs of electrons reside in orbitals that are directed toward the corners of an equilateral triangle. Proceeding in this way, it is found that NH<sub>3</sub> and BF<sub>3</sub> have the structures



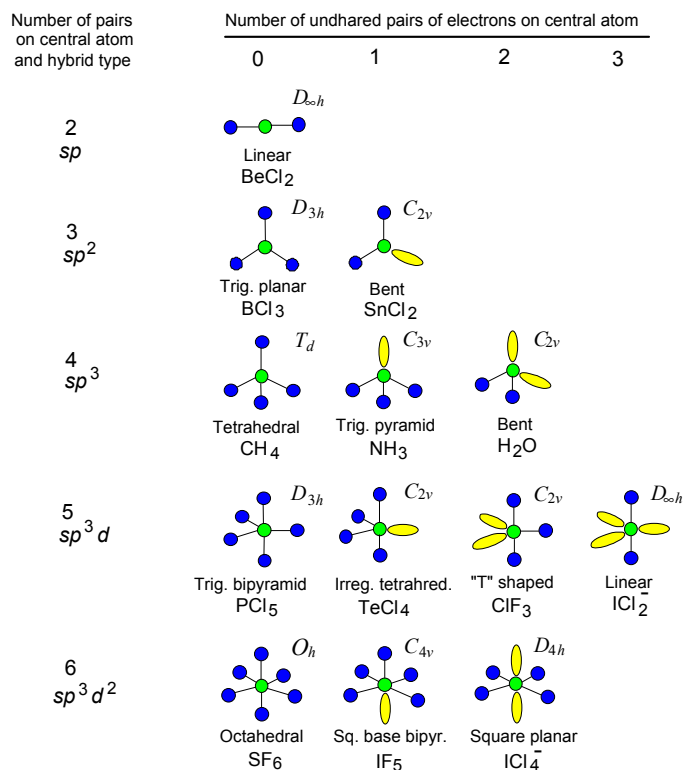
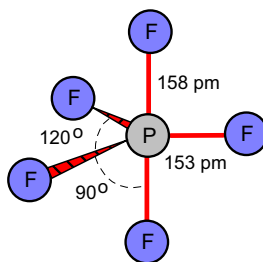


FIGURE 3.8 Geometry and symmetry (discussed in Section 3.2) of molecules.

This process could be followed for determining the structures of molecules such as  $\text{CH}_4$ ,  $\text{PF}_5$ ,  $\text{SF}_4$ ,  $\text{SF}_6$ , and many others. Figure 3.8 summarizes the common hybridization schemes and geometrical arrangements for many types of molecules. Also shown are the symmetry types (point groups), and these will be discussed later. Figure 3.8 should be studied thoroughly so that these structures become very familiar. In structures where there is  $sp$ ,  $sp^2$ ,  $sp^3$ , or  $sp^3d^2$  hybridization, all of the positions are equivalent. For a molecule such as  $\text{PF}_5$ , it is found that there are 10 electrons around the phosphorus atom, five from the P atom and one from each of five F atoms. The five pairs have minimum repulsion when they are directed toward the corners of a trigonal bipyramid so the structure can be shown as

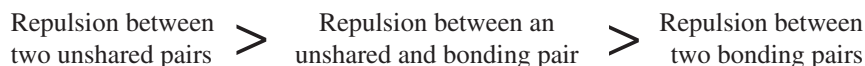


This structure is somewhat unusual in that the two bonds in axial positions are longer than those in the equatorial positions. The hybrid bond type for the trigonal bipyramid is  $sp^3d$  (or  $dsp^3$  in some cases). However, in reality this is a combination of  $dp$  (linear) and  $sp^2$  (trigonal planar) hybrids so the orbitals used in the axial and equatorial positions are not equivalent, and this is apparent from the bond lengths shown on the structure above.

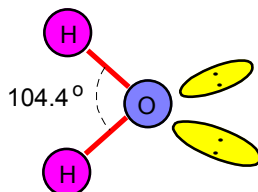
The predictions about the directions of hybrid orbitals based on minimization of repulsion are shown in Figure 3.8. Note that *any* pair of electrons requires an orbital. Note also that the molecular geometry is *not* the same as the hybrid orbital type. The oxygen atom in the  $\text{H}_2\text{O}$  molecule utilizes  $sp^3$  hybrid orbitals, but the molecule is certainly not tetrahedral! It is an angular or “V” shaped molecule. The geometry of a molecule is predicted by the hybrid orbital type only if there are no unshared pairs of electrons. Hybrid orbital type is determined by the number of electron pairs on the central atom, but the molecular geometry is determined by where the atoms are located.

### 3.1.5 Valence Shell Electron Pair Repulsion (VSEPR)

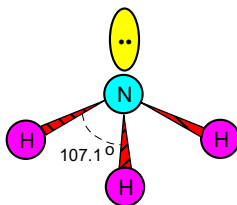
When molecules have unshared pairs of electrons (sometimes referred to as lone pairs) in addition to the bonding pairs, repulsion is somewhat different than described above. The reason is that unshared pairs are not localized between two positive nuclei as are bonding pairs. As a result, there is a difference in the repulsion between two bonding pairs and the repulsion occurring between a bonding pair and an unshared pair. Likewise, there is an even greater repulsion between two unshared pairs. Thus, with regard to repulsion,



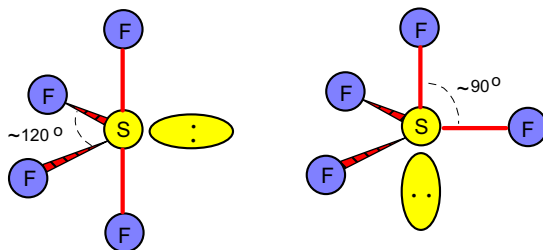
The effect of the difference in repulsion on the geometric structures of molecules is readily apparent. In methane, there are no unshared pairs so the structure is tetrahedral. However, even though  $sp^3$  hybrid orbitals are utilized by N and O in  $\text{NH}_3$  and  $\text{H}_2\text{O}$  molecules, the bond angles are not those expected for  $sp^3$  orbitals. In  $\text{H}_2\text{O}$ , the two unshared pairs of electrons repel the bonding pairs more than the bonding pairs repel each other. As expected, this repulsion causes the bond angle to be reduced from the  $109.5^\circ$  expected for a regular tetrahedral molecule. The actual H—O—H bond angle is about  $104.4^\circ$ .



For  $\text{NH}_3$ , there is only one unshared pair of electrons so its interaction with the three bonding pairs produces less reduction of the bond angle than in the case of water. The observed H—N—H angle is about  $107.1^\circ$  in accord with this expectation.

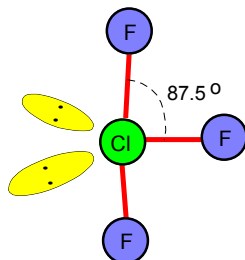


In the  $\text{SF}_4$  molecule, there are six valence electrons from the central atom and four from the four F atoms (one from each). Thus, there are 10 electrons around the central atom (five pairs) that will be directed in space toward the corners of a trigonal bipyramid. However, because there are only four F atoms with a bonding pair of electrons to each, the fifth pair of electrons must be an *unshared* pair on the sulfur atom. With there being five pairs of electrons around the central atom in  $\text{SF}_4$ , there are two possible structures that can be shown as

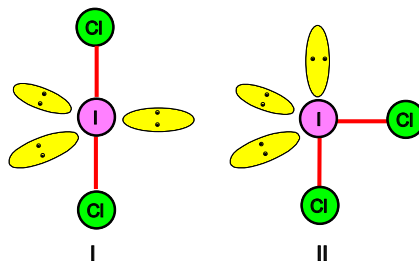


However, the correct structure is on the left with the unshared pair of electrons in an *equatorial* position. In that structure, the unshared pair of electrons is in an orbital at  $90^\circ$  from two other pairs and  $120^\circ$  from two other pairs. In the incorrect structure on the right, the unshared pair is  $90^\circ$  from three pairs and  $180^\circ$  from one pair of electrons. Although it might not seem as if there is more space in the equatorial positions, the repulsion there is less than in the axial positions. An unshared pair of electrons is not restricted to motion between two atomic centers, and it requires more space than does a shared pair. There appears to be no exception to this and all molecules in which there are five pairs of electrons around the central atom have any unshared pairs in the *equatorial* positions.

In  $\text{ClF}_3$ , which is “T” shaped, the repulsion of the unshared pairs and the bonding pairs causes the  $\text{F—Cl—F}$  bond angles to be less than  $90^\circ$ .



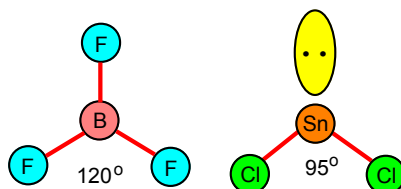
Valence shell electron pair repulsion (VSEPR) is also a useful tool to help decide which of the following structures is correct for  $\text{ICl}_2^-$ .



In structure I, the unshared pairs are all directed at  $120^\circ$  from each other and the Cl atoms lie at  $180^\circ$  from each other. In structure II, two unshared pairs lie at  $120^\circ$  from each other but the angle between them and the third unshared pair is only  $90^\circ$ . With the unshared pairs giving rise to the greatest amount of repulsion, we correctly predict that  $\text{ICl}_2^-$  would have the linear structure I.

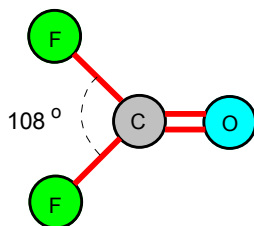
### 3.1.6 Some Subtle Influences on Bonds

Although the major features of the structure and bonding in many molecules are satisfactorily explained by the principles just described, there are other cases in which the simple approaches are inadequate. Consider the molecules  $\text{BF}_3$  and  $\text{SnCl}_2$  shown below. In each case, there are three pairs of electrons around the central atom, but the bond angles differ greatly.

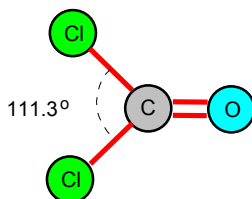


The  $\text{F—B—F}$  bond angle of  $120^\circ$  in  $\text{BF}_3$  is exactly that expected for the hybrid orbital type of  $sp^2$ . However, the three pairs of electrons around the Sn atom in  $\text{SnCl}_2$  probably do not reside in  $sp^2$  orbitals, and in fact the bond angle is closer to that expected ( $90^\circ$ ) if no hybridization of the  $p$  orbitals occurred. It is very unlikely that the effect of one unshared pair of electrons on Sn would cause a  $25^\circ$  reduction in bond angle. It should be kept in mind that hybridization schemes do not always involve simple integer ratios of orbitals.

Another interesting aspect of bonding is provided by considering the molecules  $\text{H}_2\text{C=O}$ ,  $\text{F}_2\text{C=O}$ , and  $\text{Cl}_2\text{C=O}$ . Although these molecules have structures that are similar, there are some significant differences in bond angles. The structure of  $\text{F}_2\text{C=O}$  can be shown as

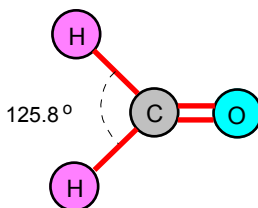


The overall structure is arrived at by considering the bonding to consist of three single bonds arising from utilization of  $sp^2$  hybrid orbitals. The double bond to the oxygen atom involves a  $\pi$ -bond that results from the nonhybridized  $p$  orbital on the carbon interacting with a filled  $p$  orbital on the oxygen atom. The deviation from expected bond angle of  $120^\circ$  is large. One might expect that repulsion between the fluorine atoms would cause the bond angle to be greater than  $120^\circ$ , but that is clearly not the case. An additional case is provided by the molecule  $\text{Cl}_2\text{C}=\text{O}$  which has the structure



The bond angle is slightly larger than in  $\text{F}_2\text{C}=\text{O}$ , which would be expected because the Cl atom is larger than the F atom. However, the difference is perhaps not as great as would be anticipated, but the bond angle is again *smaller* than the  $120^\circ$  expected if the carbon atom utilizes  $sp^2$  hybrid bonding orbitals.

If one considers the  $\text{H}_2\text{C}=\text{O}$  molecule, the situation becomes somewhat more clear. For example, the H atom is much smaller than either F or Cl so it might be expected that the bond angle would be smaller than in either  $\text{F}_2\text{C}=\text{O}$  or  $\text{Cl}_2\text{C}=\text{O}$ . However, in reality the structure is

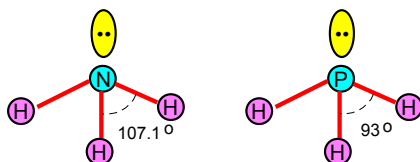


and the bond angle is *larger* than in either  $\text{F}_2\text{C}=\text{O}$  or  $\text{Cl}_2\text{C}=\text{O}$ ! Clearly, there must be some other influence that is coming into play as a factor in determining the bond angles in these molecules. The situation clears up somewhat when we consider where each of the shared pairs of electrons resides. In the case of  $\text{F}_2\text{C}=\text{O}$  and  $\text{Cl}_2\text{C}=\text{O}$ , the shared pairs are drawn closer to the F and Cl atoms, respectively, because these atoms have higher electronegativity than C. On the other hand, H has a lower electronegativity than C and the shared pairs of electrons in  $\text{H}_2\text{C}=\text{O}$  reside closer to the carbon atom, which leads to greater repulsion between the *bonding* electrons. The result is that the bond angle is *greater* in  $\text{H}_2\text{C}=\text{O}$  even though H is *smaller* than either F or Cl. Interpreting subtle aspects of molecular structure is not always as simple as looking at the hybrid orbital type on the central atom and where unshared electron pairs reside.

It should be mentioned that the irregularities in bond angles caused by VSEPR are typically only a few degrees. Qualitatively, the various hybridization schemes correctly predict the overall structure. VSEPR is, however, a very useful tool for predicting further details of molecular structure, and it will be applied many times in later chapters.

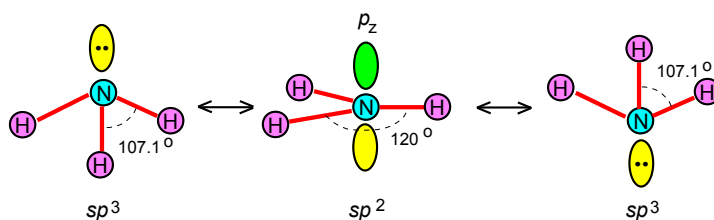
### 3.1.7 Bond Angles and Inversion

The slight deviation of the bond angles in ammonia from the expected  $109.5^\circ$  in tetrahedral molecules has been explained in terms of VSEPR. However, the bond angle in  $\text{PH}_3$  is only  $93^\circ$ .



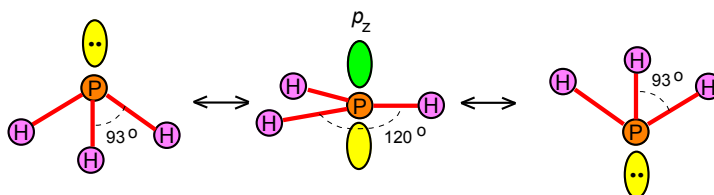
Although it may be appropriate to consider the bonding in  $\text{NH}_3$  as involving  $sp^3$  hybrid orbitals on the central atom, that is clearly not the case for  $\text{PH}_3$ . Moreover, the bond angles in  $\text{AsH}_3$ ,  $\text{SbH}_3$ , and  $\text{BiH}_3$  are only slightly greater than  $90^\circ$ . In these cases, the bond angles do not deviate greatly from the  $90^\circ$  bond angle expected if the bonding orbitals on the central atom were pure  $p$  orbitals. This may be in part the result of the fact that hybridization leads to bonding orbitals having specific directional changes but also they have different sizes. As a result, for the larger P, As, Sb, and Bi atoms, hybridization may lead to less effective overlap than if little or no hybridization occurs.

The inversion vibration of the ammonia molecule can be shown as



This vibration has a frequency of approximately  $1010 \text{ sec}^{-1}$ . Although the barrier height for inversion is  $2076 \text{ cm}^{-1}$ , the difference between the first and second vibrational states is only  $950 \text{ cm}^{-1}$  (which is equivalent to  $1.14 \text{ kJ mol}^{-1}$ ). By making use of the Boltzmann distribution law, it is possible to calculate that the second vibrational state is populated only to the extent of 0.0105, indicating that there is not a sufficient amount of thermal energy available to cause the *rapid* inversion if the molecule must pass *over* a barrier that is  $2076 \text{ cm}^{-1}$  in height. In this case, the inversion involves *quantum mechanical tunneling*, which means that the molecule passes from one structure to the other without having to pass *over* the barrier.

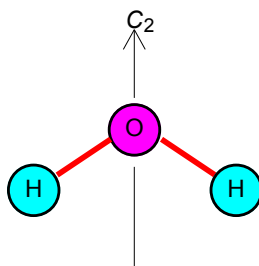
The inversion vibration of phosphine can be shown as



Although an exact value for the energy barrier is not available, it is estimated to be approximately  $140 \text{ kJ mol}^{-1}$  (Haagland, p. 228). This value is at least five times the height of the energy barrier for inversion of the  $\text{NH}_3$  molecule. As a result, inversion of  $\text{PH}_3$  does not occur rapidly. This phenomenon is not completely understood, but it may well indicate that the planar transition state of the  $\text{PH}_3$  molecule does not involve  $sp^2$  hybrid bonds, but rather that  $d$  orbital participation is involved. The difference between  $\text{NH}_3$  and  $\text{PH}_3$  shows that one cannot infer that behavior of a compound containing a first row element is a good predictor of behavior of a similar compound containing a heavier atom in the same group.

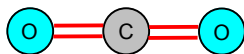
### 3.2 SYMMETRY

One of the most efficient ways to describe the spatial arrangement of atoms in a molecule is to specify its *symmetry*, which allows a symbol to be used to specify a great deal of information succinctly. In examining the structure of a molecule from the standpoint of symmetry, lines, planes, and points are identified that are related to the structure in particular ways. Consider the  $\text{H}_2\text{O}$  molecule

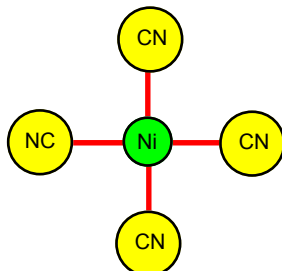


It is apparent that there is a line through the O atom that bisects the H—O—H angle. Rotation of the molecule around this line by  $180^\circ$  leaves the structure unchanged (said to be an indistinguishable orientation). This line is referred to as a *symmetry element*, an axis of rotation (more precisely, a *proper rotation axis*). The process of rotating the molecule is a *symmetry operation*. The mathematical rules governing symmetry operations and their combinations and relationships involve group theory. For more details on the application of group theory, the references at the end of this chapter should be consulted. The purpose here is to identify the symmetry elements and arrive at symmetry designations for molecules. The various symmetry elements are as follows.

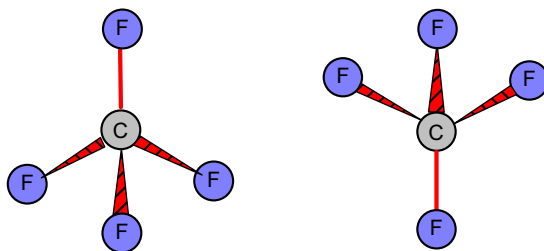
1. A *center of symmetry* or an *inversion center* (*i*). A molecule possesses a *center of symmetry* if inversion of each atom through this center results in an identical arrangement of atoms. For example,  $\text{CO}_2$  has a center of symmetry,



which is at the center of the carbon atom. In  $\text{Ni}(\text{CN})_4^{2-}$  there is also a center of symmetry.

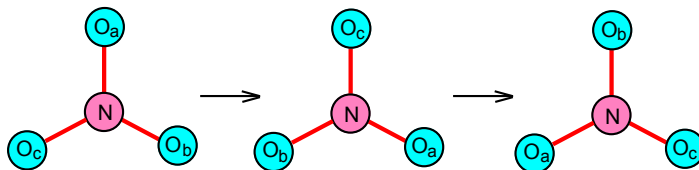


The Ni atom is at the center of the structure and inversion of each atom through that point gives exactly the same arrangement shown above. However, in the tetrahedral  $\text{CF}_4$  molecule, inversion of each atom through the C atom gives a different result.



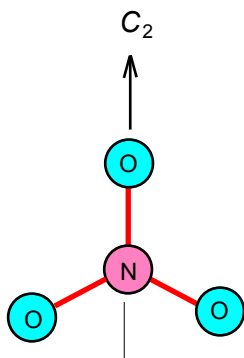
The geometric center of the  $\text{CF}_4$  molecule is, therefore, not a center of symmetry.

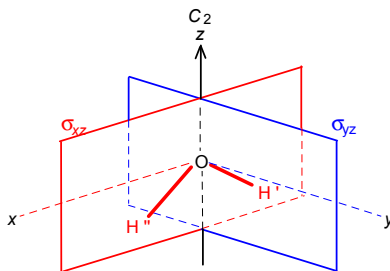
2. The *proper rotation axis* ( $C_n$ ). If a molecule can be rotated around an imaginary axis to produce an equivalent orientation, the molecule possesses a *proper rotation axis*. The line in the  $\text{H}_2\text{O}$  structure shown above is such a line. Consider the planar  $\text{NO}_3^-$  ion.



In the view shown, the axis projecting out of the page (the  $z$ -axis) is a line around which rotation by  $120^\circ$  gives an indistinguishable orientation. In this case, because the rotations producing indistinguishable orientations are  $120^\circ$  or  $360^\circ/3$ , the rotation axis is referred to as a threefold or  $C_3$  axis. Three such rotations (two of which are shown above) return the molecule to its original orientation.

However, in the case of  $\text{NO}_3^-$ , there are also twofold axes that lie along each N–O bond.

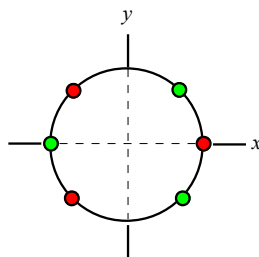




**FIGURE 3.9** The water molecule showing two mirror planes. The intersection of these two planes generates a  $C_2$  axis. The plane outlined in red bisects all three atoms whereas that outlined in blue bisects the oxygen atom and leaves one hydrogen atom on either side.

Rotation of the molecule around the axis shown leaves the position of one oxygen atom unchanged and the other two interchanged. Thus, this axis is a  $C_2$  axis because rotation by  $180^\circ$  or  $360^\circ/2$  produces an identical orientation. Although there are three  $C_2$  axes, the  $C_3$  axis is designated as the *principal axis*. The principal axis is designated as the one of highest fold rotation. This is the customary way of assigning the  $z$ -axis in setting up an internal coordinate system for the molecule.

3. *The mirror plane (plane of symmetry) ( $\sigma$ )*. If a molecule has a plane that divides it into two halves that are mirror images, the plane is called a *mirror plane (plane of symmetry)*. Consider the  $H_2O$  molecule as shown in Figure 3.9. The O–H bonds lie in the  $yz$ -plane. Reflection of the hydrogen atoms through the  $xz$ -plane interchanges the locations of  $H'$  and  $H''$ . Reflection through the  $yz$ -plane interchanges the “halves” of the hydrogen atoms lying on the  $yz$ -plane. It is should be apparent that the intersection of the two planes generates a line (the  $z$ -axis), which is therefore a  $C_2$  axis. As a result of the  $z$ -axis being the principal axis, both of the planes shown are *vertical planes* ( $\sigma_v$ ).
4. *Improper rotation axis ( $S_n$ )*. An *improper rotation axis* is one about which rotation followed by reflection of each atom through a plane *perpendicular* to the rotation axis produces an identical orientation. Thus, the symbol  $S_6$  means to rotate the structure clockwise by  $60^\circ$  ( $360^\circ/6$ ) and reflect each atom through a plane perpendicular to the axis of rotation. This can be illustrated as follows. Consider the points lying on the coordinate system shown below with the  $z$ -axis projecting out of the page. A green circle indicates a point lying below the  $xy$ -plane (the plane of the page) whereas a red circle indicates a point lying above that plane.



It can be seen that the line through the origin pointing directly out of the page, the  $z$ -axis, is a  $C_3$  axis. However, rotation of the structure around that axis by  $60^\circ$  followed by reflection through the  $xy$ -plane (the page) moves the objects to exactly the same positions shown in the figure. Therefore, the  $z$ -axis is an  $S_6$  axis. This structure is that exhibited by the cyclohexane molecule in the “chair” configuration.



Consider a tetrahedral structure such as that shown in Figure 3.10(a). Rotation by  $180^\circ$  around the  $x$ -,  $y$ -, or  $z$ -axis leaves the structure unchanged. Therefore, these axes are  $C_2$  axes. However, if the structure is rotated by  $90^\circ$  around the  $z$ -axis, the result is shown in Figure 3.10(b).

Another symmetry element is illustrated by Figure 3.11. In this case, the molecule in its original orientation (Figure 3.11(a)) is rotated around the  $z$ -axis by  $90^\circ$  after which each atom is reflected through the  $xy$ -plane. This rotation moves the atoms to the positions shown in Figure 3.11(b), which is identical to that of the original. Therefore, the  $z$ -axis is an improper rotation axis,  $S_4$ . In a similar way, it is easy to show that the  $z$ - and  $y$ -axis are also  $S_4$  axes so a molecule such as  $CH_4$  possesses three such axes.



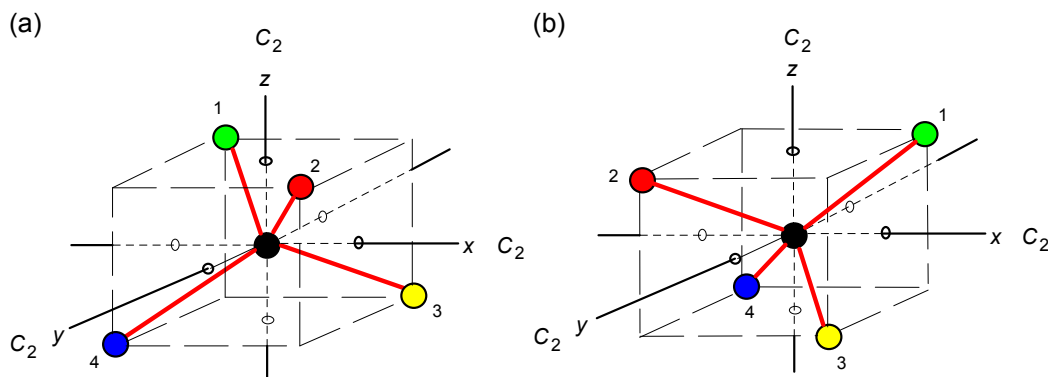


FIGURE 3.10 (a) A tetrahedral molecule. (b) The same molecule after rotation by  $90^\circ$  around the  $z$ -axis.

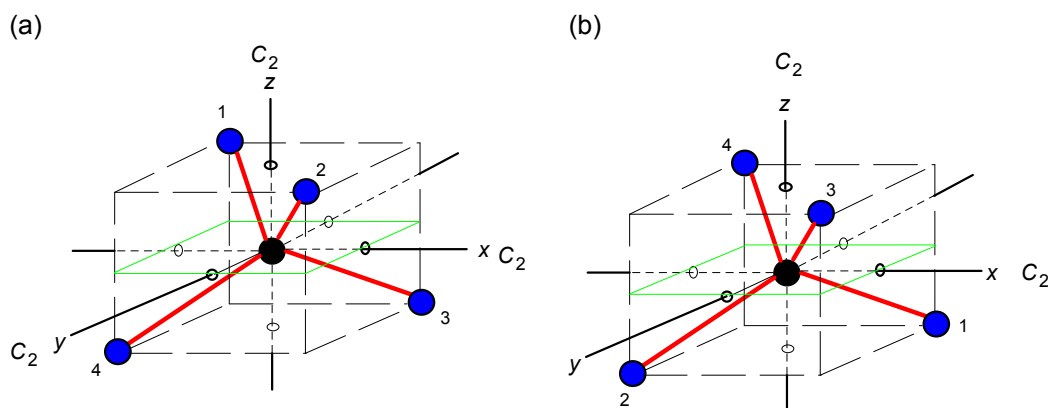
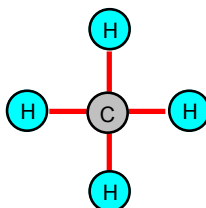


FIGURE 3.11 A tetrahedral molecule before rotation (a) and after rotation (b) by  $90^\circ$  around the  $z$ -axis and reflection of each atom through the  $xy$ -plane. After this operation, the structures become indistinguishable.

After deciding what symmetry elements are present, a molecule can be assigned to a symmetry category known as a *point group*, which is based on the collection of symmetry elements the molecule possesses. Table 3.3 shows a summary of some of the most important point groups, lists the symmetry elements present, and shows some examples of each type of structure.

It should be emphasized that in order to determine what symmetry elements are present, it is necessary to have a good representation of the structure. A diagram showing the correct perspective should be drawn before trying to pick out symmetry elements. For example, if the structure of  $\text{CH}_4$  is drawn as



(which is incorrect) there *appears* to be a  $C_4$  axis through the carbon atom perpendicular to the page. From the incorrect structure, it is also going to be very hard to visualize that each C–H bond is actually a  $C_3$  axis. Structures should always be drawn with the correct geometry indicated by the examples as shown in Figure 3.8.

**TABLE 3.3** Common Point Groups and Their Symmetry Elements

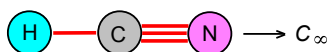
Point Group	Structure	Symmetry Elements	Examples
$C_1$	—	None	CHFClBr
$C_s$	—	One plane	ONCl, OSCl <sub>2</sub>
$C_2$	—	One $C_2$ axis	H <sub>2</sub> O <sub>2</sub>
$C_{2v}$	AB <sub>2</sub> bent or XAB <sub>2</sub> planar	One $C_2$ axis and two $\sigma_v$ at 90°	H <sub>2</sub> O, SO <sub>2</sub> NO <sub>2</sub> , H <sub>2</sub> CO
$C_{3v}$	AB <sub>3</sub> pyramidal	One $C_3$ axis and three $\sigma_v$ planes	NH <sub>3</sub> , PH <sub>3</sub> , CHCl <sub>3</sub>
$C_{nv}$	—	One $C_n$ axis and $n$ $\sigma_v$ planes	BrF <sub>5</sub> ( $C_{4v}$ )
$C_{\infty v}$	ABC linear	One $C_\infty$ axis and $\infty$ $\sigma_v$ planes	HCN, SCO, OCN <sup>-</sup> , SCN <sup>-</sup>
$D_{2h}$	Planar	Three $C_2$ axes, two $\sigma_v$ planes, one $\sigma_h$ plane, and center of symmetry	C <sub>2</sub> H <sub>4</sub> , N <sub>2</sub> O <sub>4</sub>
$D_{3h}$	AB <sub>3</sub> planar	One $C_3$ axis, three $C_2$ axes, three $\sigma_v$ and one $\sigma_h$ plane	BF <sub>3</sub> , CO <sub>3</sub> <sup>2-</sup> , NO <sub>3</sub> <sup>-</sup> , SO <sub>3</sub>
$D_{4h}$	AB <sub>4</sub> planar	One $C_4$ and four $C_2$ axes, one $\sigma_h$ and four $\sigma_v$ planes, and center of symmetry	XeF <sub>4</sub> , PtCl <sub>4</sub> <sup>2-</sup>
$D_{\infty h}$	AB <sub>2</sub> linear	One $C_\infty$ axis, $\infty$ $C_2$ axes, $\infty$ $\sigma_v$ and one $\sigma_h$ planes, and center of symmetry	CO <sub>2</sub> , NO <sub>2</sub> <sup>+</sup> CS <sub>2</sub>
$T_d$	AB <sub>4</sub>	Four $C_3$ and three $C_2$ axes, six $\sigma_v$ planes, and three $S_4$ axes	CH <sub>4</sub> , P <sub>4</sub> MnO <sub>4</sub> <sup>-</sup> , SO <sub>4</sub> <sup>2-</sup>
$O_h$	AB <sub>6</sub> octahedral	Three $C_4$ , four $C_3$ , six $C_2$ , four $S_6$ , and three $S_4$ axes, nine $\sigma_v$ planes center of symmetry	SF <sub>6</sub> , Cr(CO) <sub>6</sub> , PF <sub>6</sub> <sup>-</sup>
$I_h$	Icosahedral	6 $C_5$ , 10 $C_3$ , and 15 $C_2$ axes, 15 planes, 20 $S_6$ axes	B <sub>12</sub> , B <sub>12</sub> H <sub>12</sub> <sup>2-</sup>

In order to illustrate how symmetry concepts can be applied to molecules, several examples will be considered.

### Example 3.1

#### HCN

This molecule has a structure

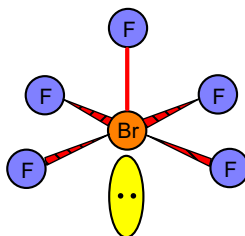


The line through all three nuclei is a  $C_\infty$  axis because rotation around it by *any* angle, no matter how small, gives the same orientation. There are an infinite number of planes ( $\sigma_v$ ) that intersect along this  $C$  axis. A molecule having a  $C_n$  axis and  $n$  vertical planes intersecting along it is called a  $C_{nv}$  structure. Therefore, the point group of the HCN molecule is  $C_{\infty v}$ .

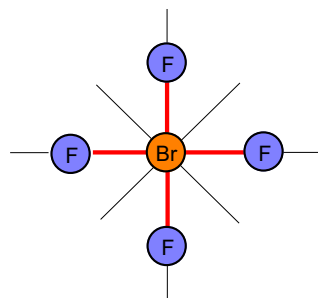
### Example 3.2

#### BrF<sub>5</sub>

The structure of the BrF<sub>5</sub> molecule is



A line through the fluorine atom in the axial position and the bromine atom is a  $C_4$  axis, the  $z$ -axis. There are four vertical planes that intersect along that axis. Looking down the  $C_4$  axis, the planes appear edge on as shown in this structure.

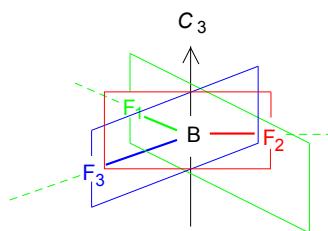


Therefore, in view of the fact that the molecule has a  $C_4$  axis and four vertical planes, it belongs to the point group  $C_{4v}$ .

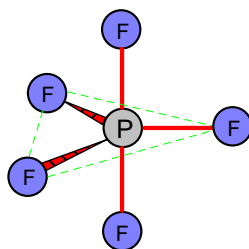
### Example 3.3

#### $\text{BF}_3$

The structure of  $\text{BF}_3$  is a trigonal plane as shown below with the rotation axes and planes of symmetry as indicated.



For  $\text{BF}_3$ , there is a  $C_3$  axis as shown in the structure and the molecule lies in a horizontal plane, designated as  $\sigma_h$ . There are also three vertical planes. Each B–F bond also defines a  $C_2$  axis where the  $\sigma_h$  intersects each  $\sigma_v$ . Therefore, the trigonal planar  $\text{BF}_3$  molecule has one  $C_3$  axis, one  $\sigma_h$ , three  $\sigma_v$ , and three  $C_2$  axes. A molecule that has a  $C_n$  axis and  $nC_2$  axes perpendicular to it resulting from the horizontal plane intersecting the three vertical planes belongs to the  $D_{nh}$  point group. Therefore,  $D_{3h}$  is the appropriate point group for  $\text{BF}_3$ . Note that a structure such as that of  $\text{PF}_5$ ,

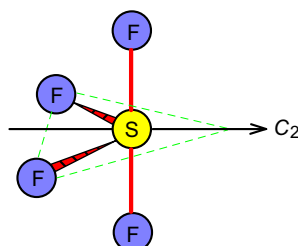


also has one  $C_3$  axis, three  $\sigma_v$ , one  $\sigma_h$ , and three  $C_2$  axes. Therefore, it also belongs to the point group  $D_{3h}$ .

### Example 3.4

#### $\text{SF}_4$

The structure of the  $\text{SF}_4$  molecule can be shown as

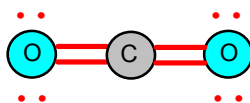


The structure possesses one  $C_2$  axis (as shown), and there are two mirror planes that intersect along it so this molecule belongs to the  $C_{2v}$  point group.

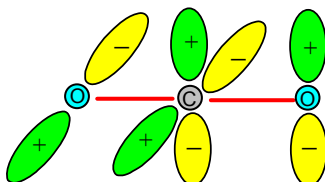
The reader should now go back to Figure 3.8 and examine the structures shown and their corresponding point groups. When making drawings for molecules, make them so that the symmetry elements can be identified and the point groups assigned by referring to Table 3.2. In working through this book, a large number of molecular structures will be encountered, and the point groups are indicated for many of them. With practice, the visualization of molecular structures and recognition of the point groups to which they belong become routine. It also makes it possible to recognize the spatial orientations of the bonds, which is often important in predicting reactive sites in molecules.

### 3.3 RESONANCE

Although structures for many chemical species can be described using only single bonds, this is not possible for many others. For the relatively simple  $\text{CO}_2$  molecule, the structure can be determined as follows. The carbon atom has four valence shell electrons and each oxygen atom has six, so 16 valence shell electrons must be distributed to provide an octet around each atom. For three atoms, eight electrons around each would require a total of 24 electrons if there were no shared electrons. Because there are only 16 electrons available, eight must be shared so they contribute to the octet of more than one atom. Eight shared electrons would constitute four bonds so the structure of  $\text{CO}_2$  can be shown as follows.

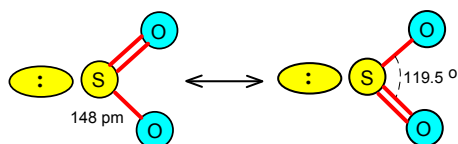


Repulsion between the four shared pairs will be minimized when the bond angle is  $180^\circ$ . The  $\text{CO}_2$  molecule is linear and has a center of symmetry so it belongs to the point group  $D_{\infty h}$ . The double bonds between carbon and oxygen atoms arise from the utilization of  $sp$  hybrid orbitals on the carbon atom overlapping with  $p$  orbitals on the oxygen atoms to give a  $\sigma$  bond and the overlapping of  $p$  orbitals on the carbon with  $p$  orbitals on the oxygen atoms to give two  $\pi$  bonds. These bonds can be illustrated as shown here.



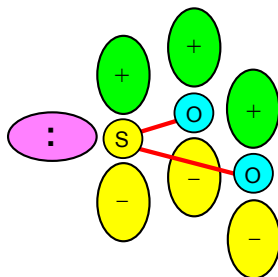
Note that the  $p_z$  orbital on the carbon atom is utilized in the  $sp$  hybrids because the  $z$ -axis lies along the  $\text{O}-\text{C}-\text{O}$  axis. Numerous other species, e.g.,  $\text{NO}_2^+$ ,  $\text{OCN}^-$ , and  $\text{SCN}^-$ , have 16 valence electrons and linear structures.

When sulfur burns in air, the product is  $\text{SO}_2$ . The three atoms have a total of 18 valence shell electrons, and by following the procedure described above, it found that six electrons must be shared. This means that there will be three bonds to the central atom. However, for it to have an octet of electrons, it must also have an unshared pair of electrons. The structures that show these features are



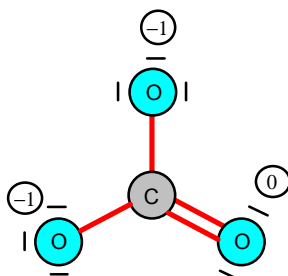
The significance of the double headed arrow is to indicate that the molecule actually has a structure that is a composite of the two structures shown. In this case, the structures are identical, and they contribute equally, but this is not always so. The term *resonance* is used to describe structures that differ only in the placement of electrons so the structures above are referred to as the *resonance structures* for  $\text{SO}_2$ .

Because of the angular structure of the molecule, it appears that the sulfur atom makes use of  $sp^2$  hybrid orbitals. The remaining  $p$  orbital is utilized in forming a  $\pi$  bond that results from overlap with  $p$  orbitals on the oxygen atoms. This can be shown as

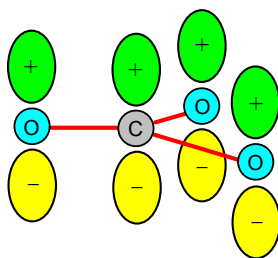


Note that in this case each bond between S and O is the average of one single and one double bond, which gives a bond order of 1.5. In accord with this conclusion is the fact that the bond lengths are equal, and they are intermediate between the normal lengths of S—O and S=O bonds. Other diatomic species that exhibit this type of structure are  $\text{NO}_2$ ,  $\text{NO}_2^-$ , and  $\text{O}_3$ .

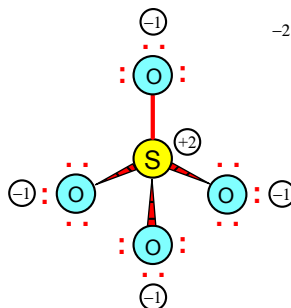
In determining the structure of  $\text{CO}_3^{2-}$ , it is necessary to arrange 24 electrons (4 from C and 6 from each O, and 2 from the  $-2$  charge). Complete octets for 4 atoms would require 32 electrons so 8 electrons must be shared. Four bonds are expected, which in this case means there must be two single bonds and one double bond. The structure consistent with this distribution of electrons is



The double bond that arises from overlap of the carbon  $2p$  orbital with one of the oxygen  $2p$  orbitals can be in any of the three positions. The overall structure is consistent with an  $sp^2$  hybridized carbon atom with one  $p$  orbital remaining. In this case, there are three equivalent structures that can be drawn with the double bond “smeared out” over the three oxygen atoms giving a bond order of  $4/3$  to each oxygen. The structure is planar with three  $\sigma$  bonds and one  $\pi$  bond, and the point group is  $D_{3h}$ . It is apparent that the double bond could be in any of three positions so there are three resonance structures that can be combined to give a  $\pi$  bond that extends over the entire structure as follows.



In  $\text{SO}_4^{2-}$ , there are 32 valence shell electrons. Five atoms having octets would require 40 electrons, so in this case, there are eight shared electrons in four bonds, one to each oxygen. The structure, therefore, can be shown as



which is tetrahedral ( $T_d$  point group). This is also the structure of  $\text{PO}_4^{3-}$ ,  $\text{MnO}_4^-$ ,  $\text{ClO}_4^-$ , and other species. However, as will be described, the structure for  $\text{SO}_4^{2-}$  is not as simple as indicated here.

There are characteristics or rules that apply to the drawing of resonance structures. These can be summarized as follows:

1. The relative positions of the atoms must be the same in all structures.
2. There must be the same number of unpaired electrons (if any) in all structures.
3. Resonance structures that maximize the number of electrons used in bonding will likely contribute a greater fraction to the true structure.
4. Resonance structures should preserve an octet of electrons around each atom if possible.
5. These rules apply concerning formal charges.
  - (a) Like formal charges on adjacent atoms leads to instability.
  - (b) Atoms of high electronegativity should have negative or zero formal charges.
  - (c) Formal charges should not be large in magnitude.
  - (d) The sum of the formal charges on all the atoms must equal the overall charge on the species.

Certainly there are cases where it is not possible to obey Rule 4. For example in  $\text{SbF}_5$  there must be one pair of electrons forming a bond to each fluorine atom. Therefore, antimony has 10 electrons around it. Also, in  $\text{BCl}_3$ , there are only six electrons surrounding the boron atom.

The two resonance structures shown for  $\text{SO}_2$  contribute equally to the actual structure as did the three resonance structures shown for  $\text{CO}_3^{2-}$ . However, resonance structures do not always contribute equally, and it necessary to have some way of estimating their relative contributions. One of the most useful tools for this is provided by the concept of *formal charges*.

As mentioned above, formal charges provide a useful tool for estimating contributions from resonance structures. The formal charge on an atom gives a comparison of the electron density around an atom in a molecule to that surrounding the uncombined atom. It must be emphasized that the *formal charge* is in no way an actual electrostatic charge, but it is a way of keeping track of electrons, an electron “accounting” procedure. The formal charge is obtained by finding the difference between the number of electrons an atom has surrounding it in a molecule and the number of valence electrons the atom normally has. The number of electrons surrounding an atom in the molecule is obtained as follows.

1. Count any unshared pairs on the atom as belonging to that atom.
2. Consider half of each bonding pair as belonging to each of the bonded atoms.
3. The sum obtained from steps 1 and 2 gives the number of electrons assigned to the atom in question in the molecule ( $e_m$ ).
4. The formal charge,  $f$ , is given by

$$f = e_v - e_m \quad (3.14)$$

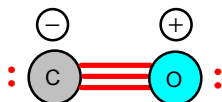
where  $e_v$  is the usual number of valence shell electrons.

5. As previously stated, the sum of the formal charges on all the atoms must equal the overall charge on the species.

For carbon monoxide, the usual valence bond structure is shown as follows.

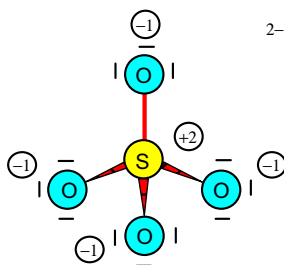


In this structure, there is an octet of electrons around each atom. Note that the triple bond is consistent with the bond order of three predicted earlier from the MO diagram. In calculating the formal charge on the oxygen atom, it can be seen that the unshared pair contributes two electrons and there is one from each of the three bonding pairs. This gives a total of five electrons around the oxygen atom. Because the oxygen atom has six valence shell electrons, the formal charge is  $6 - 5 = +1$ . In a similar way, it can be seen that the carbon atom *appears* to have five electrons surrounding it in the structure above so its formal charge is  $4 - 5 = -1$ . It should also be noted that the sum of the formal charges on the atoms is zero, which is correct for the CO molecule. Therefore, the formal charges are shown as

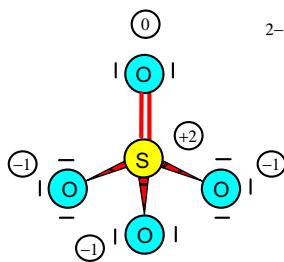


where the charges are circled to distinguish them from ionic charges. Based on the greatly differing electronegativities, one might think that CO should be very polar. In fact, however, the dipole moment of CO is only about 0.12 D with carbon lying at the *negative* end and oxygen at the *positive* end. The formal charges shown in the structure above predict this aspect of the chemical characteristics of the CO molecule. When CO forms complexes with metals, it is the carbon end of the molecule that functions as an electron pair donor as would be expected on the basis of formal charges.

The sulfate ion provides another example of the value of formal charges in explaining resonance. When the structure is drawn as if the S–O bonds were single bonds, it is seen that there is a +2 formal charge on the sulfur atom.

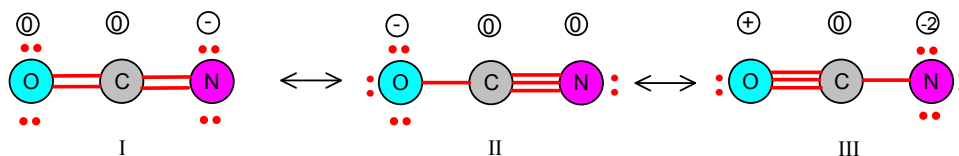


This is rather high for an atom that has an electronegativity of 2.6. It is possible to draw resonance structures that give a more equitable distribution of electrons and one of them is shown below.



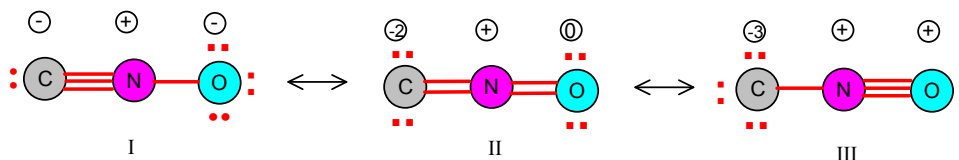
There would be four such structures that result from placing the double bond in the four available positions. The inclusion of these structures indicates that the S–O bonds should have a bond order that is greater than 1, and that the bonds should be shorter than expected for a single bond, which is in fact the case. Formal charges provide a useful tool for elucidating some of the subtle aspects of structures.

The application of formal charges with regard to resonance can now be illustrated by considering the structures of  $\text{OCN}^-$  (the cyanate ion) and  $\text{CNO}^-$  (the fulminate ion). Both of these ions are triatomic species having 16 valence electrons and belong to the  $C_{\infty v}$  point group. At first glance, it might appear that there might not be much difference in the stability of these two species. If the resonance structures are drawn for  $\text{OCN}^-$  it is seen that the three most important structures are the following



An examination of structure III shows that it is not likely to contribute very much to the true structure of  $\text{OCN}^-$  because it violates some of the principles related to resonance structures. Structure III has a positive formal charge on the atom of highest electronegativity, and it has higher overall formal charges. Both of the other structures contribute significantly to the structure  $\text{OCN}^-$ , and a reasonable estimate is that their contributions should be approximately equal.

Three resonance structures that can be drawn for  $\text{CNO}^-$  are shown as follows.



It should be readily apparent that structure III would not contribute to the true structure (see rules above). Structure II also has the unlikely situation of having a formal charge of -2 on the *least* electronegative atom, C. Accordingly, it will not

contribute significantly to the true structure. Finally, structure I also has a negative formal charge on the least electronegative atom, C, and a positive formal charge on the more electronegative atom, N. Therefore, even this structure is not likely to correspond to a species of high stability. As a result of this analysis, it can be seen that  $\text{OCN}^-$  (cyanate) has two resonance structures that obey the rules closely enough to contribute significantly to the structure. In contrast,  $\text{CNO}^-$  (fulminate) has only one structure that is at all plausible, and even it has flaws. Consequently, on the basis of structures, great differences would be predicted between the chemistry of cyanates and fulminates. In fact, cyanates are generally stable whereas fulminates have been used as detonators, properties that are in agreement with our analysis of their resonance structures. From the comparison of  $\text{OCN}^-$  and  $\text{CNO}^-$ , it can be seen that chemical behavior can often be interpreted by making use of principles of structure and bonding.

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## PROBLEMS

- For each of the following, draw the MO energy level diagram, determine the bond order, and tell whether the species is paramagnetic or diamagnetic.
  - $\text{F}_2^+$
  - $\text{CN}^-$
  - $\text{C}_2^{2-}$
  - $\text{CN}^+$
  - BN
- Construct the MO diagram for LiH and discuss its major differences from those of  $\text{Li}_2$  and  $\text{H}_2$ .
- For each of the following, draw a MO energy level diagram, place the electrons in the orbitals appropriately, tell whether the species is paramagnetic or diamagnetic, and give the bond order.
  - $\text{NO}^-$
  - $\text{C}_2^-$
  - $\text{BN}^{2-}$
  - $\text{O}_2^+$
- The CN *molecule* has an absorption band at  $9000\text{ cm}^{-1}$  (near infrared) whereas the  $\text{CN}^-$  *ion* does not. Draw the MO diagrams for these species and explain this difference.
- What is the percent ionic character of the CIF bond?
  - If the CIF bond length is 212 pm, what would be the dipole moment of the molecule?
- For a diatomic molecule XY,  $\psi_{\text{molecule}} = \psi_{\text{covalent}} + 0.50\psi_{\text{ionic}}$ . Calculate the percent ionic character of the X–Y bond. If the X–Y bond length is 150 pm, what is the approximate dipole moment of XY?
- Suppose two elements A and B form a diatomic molecule AB.
  - If the difference in electronegativities of A and B is 1.2, what will be the ionic contribution to the A–B bond?
  - If the A–B bond length is 150 pm, what will be the dipole moment of AB?
  - If the A–A and B–B bond strengths are 240 and 425  $\text{kJ mole}^{-1}$ , respectively, what should be the strength of the A–B bond?
- Draw structures for the following showing correct geometry and all valence shell electrons.
  - $\text{XeF}_2$
  - $\text{SbF}_4^+$
  - $\text{TeCl}_4$



- (d)  $\text{ClO}_2^-$   
 (e)  $\text{CN}_2^{2-}$   
 (f)  $\text{ICl}_2^+$
- For each of the species shown in Question 8, list all symmetry elements and assign the symmetry type (point group).
  - Suppose a molecule contains one atom of phosphorus, one atom of nitrogen, and one atom of oxygen.
    - What would be the arrangement of atoms in this molecule? For this arrangement of atoms, draw the possible resonance structures.
    - Which of the structures you have shown is the least stable? Explain your answer.
  - Consider the  $-1$  ion made up of one atom each of carbon, oxygen, and phosphorus. What is the arrangement of atoms in the structure of this ion? Draw possible resonance structures for this ion. Which is the least stable structure? Explain your answer.
  - Although the  $\text{CO}_2$  molecule is nonpolar,  $\text{OCS}$  is polar. Draw resonance structures for the molecules and explain this difference.
  - The P–O and Si–O single bond lengths are 175 and 177 pm, respectively. However, in the  $\text{PO}_4^{3-}$  and  $\text{SiO}_4^{4-}$  ions, the P–O and Si–O bond lengths are 154 and 161 pm, respectively.
    - Why are the bond lengths in the ions not equal to the single bond lengths?
    - Why is the P–O bond shortened more (as compared to the single bond length) than in the Si–O bond?
  - The bond angles  $117^\circ$ ,  $134^\circ$ , and  $180^\circ$  belong to the species  $\text{NO}_2^+$ ,  $\text{NO}_2^-$ , and  $\text{NO}_2$ , but not necessarily in that order. Match the species with their bond angles and explain your answer.
  - Draw structures for the following showing correct geometry. List all symmetry elements and assign the symmetry type (point group) for each.
    - $\text{PH}_3$
    - $\text{SO}_3$
    - $\text{ClO}_4^-$
    - $\text{ClF}_3$
    - $\text{ICl}_4^-$
    - $\text{OPCl}_3$
  - For each of the following, draw the structure showing the correct geometry and all valence shell electrons. List all symmetry elements and assign the symmetry type (point group) for each.
    - $\text{PCl}_5$
    - $\text{SnCl}_2$
    - $\text{XeF}_4$
    - $\text{SO}_3^{2-}$
    - $\text{XeO}_3$
  - Match each of the following species,  $\text{ClF}_3$ ,  $\text{HCN}$ ,  $\text{SiF}_4$ ,  $\text{SO}_3^{2-}$ ,  $\text{NO}_2^+$ , with the symmetry element that describes it.
    - Has three  $C_2$  rotation axes
    - Has a center of symmetry
    - Has  $C_{2v}$  symmetry
    - Has  $C_{\infty v}$  symmetry
    - Has only one  $C_3$  axis
  - Use the principles described in Section 3.3 to determine the structure of  $\text{N}_2\text{O}$  and explain your answer.
  - Would resonance structures involving double bonds be more, less, or of the same importance in  $\text{PO}_4^{3-}$  compared to  $\text{SO}_4^{2-}$ ? Justify your answer.
  - It has been suggested that the difference in the inversion of  $\text{NH}_3$  and  $\text{PH}_3$  may be due in part to the greater repulsion between the hydrogen atoms in  $\text{NH}_3$ . Based on the bond angles and bond lengths of the molecules, does this seem reasonable? Why or why not?
  - By means of resonance structures and their contributions, explain the difference in dipole moments (shown in Table 3.2) of the  $\text{SO}_2$  and  $\text{SeO}_2$  molecules.

# Ionic Bonding, Crystals, and Intermolecular Forces

Although molecules are held together by bonds that are predominantly covalent, many substances are made up of ions that are arranged in a crystal lattice. These materials are held together in the solid state by forces that are essentially electrostatic in character. In some cases, the forces arise from the transfer of electrons between atoms to produce ionic materials. However, in most cases the ions are somewhat polarizable (especially anions) so the ions have distorted structures that represent some degree of electron sharing. As a result, many of the forces in crystals that are normally considered to be “ionic” may be appreciably less than completely ionic. This fact should be kept in mind as the principles of ionic bonding are discussed.

In order to predict and correlate the properties of inorganic compounds, it is essential to realize that there are other forces that are important in chemistry. These are forces that exist *between* molecules that can have a great effect on physical and chemical properties. If the molecules are polar, there will be an electrostatic attraction between them as a result of dipole–dipole forces. The hydrogen bond is a special type of this kind of interaction. Additionally, even though molecules may not be polar, there are weak forces that hold the molecules together in the solid (e.g., solid CO<sub>2</sub>) and liquid states. An understanding of these intermolecular forces is necessary in order to understand the chemical and physical behavior of all materials regardless of whether they are of inorganic, organic, or biological origin.

## 4.1 IONIC BONDS

When atoms having greatly differing electronegativities combine, electron transfer occurs and ions are formed. The electrostatic attraction between ions is the essence of ionic bonding. However, the transfer may not be complete so there is a continuum of bond character from complete covalent (as in H<sub>2</sub>) to almost completely ionic (as in LiF).

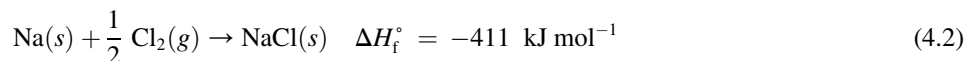
### 4.1.1 Energetics of the Ionic Bond

When ions are formed as the result of electron transfer between atoms, the charged species produced interact according to Coulomb’s law with a force,  $F$ , given by

$$F = \frac{q_1 q_2}{\epsilon r^2} \quad (4.1)$$

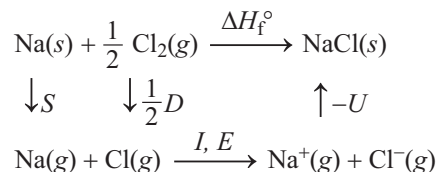
where  $q_1$  and  $q_2$  are the charges,  $r$  is the distance of separation, and  $\epsilon$  is the dielectric constant of the medium separating the ions (1 for a vacuum). Although the arrangement of ions in the solid lattice is of considerable importance, the primary concern at this point is a description of the energy changes that accompany the formation of ionic bonds.

Sodium chloride is formed from the elements in their standard states with a heat of formation of  $-411 \text{ kJ mol}^{-1}$ .



This process can be represented as taking place in a series of steps, and these steps have associated with them enthalpy changes that are known. After the steps have been written, the application of Hess’ law provides a convenient way to obtain the enthalpy for the overall process. The various steps in the formation of crystalline NaCl from the elements can be summarized in a thermochemical cycle known as a *Born–Haber cycle* (shown below). In this cycle, the heat change

for the formation of  $\text{NaCl}(s)$  is the same regardless of which way the reaction goes between the starting and ending points. The reason for this is that the *enthalpy* change is a function of initial and final states, not the pathway between them.



In the Born–Haber cycle for the formation of  $\text{NaCl}$ ,  $S$  is the sublimation enthalpy of  $\text{Na}$ ,  $D$  is the dissociation enthalpy of  $\text{Cl}_2$ ,  $I$  is the ionization potential of  $\text{Na}$ ,  $E$  is the electron addition enthalpy of  $\text{Cl}$  (which is the negative of the electron affinity), and  $U$  is the lattice energy. It can be seen from the Born–Haber cycle that the lattice energy corresponds to the energy required to *separate* a mole of crystal into the gaseous ions and *forming* the crystal from the ions represents  $-U$ .

In some cases, the energy to be determined is the lattice energy,  $U$ , because the enthalpy change for the formation of  $\text{NaCl}(s)$  is the same for both pathways. Therefore,

$$\Delta H_f^\circ = S + \frac{1}{2} D + I + E - U \quad (4.3)$$

and solving for  $U$  gives

$$U = S + \frac{1}{2} D + I + E - \Delta H_f^\circ \quad (4.4)$$

In the case of  $\text{NaCl}$ , substituting for the quantities shown on the right-hand side of the equation leads to  $U(\text{kJ mol}^{-1}) = 109 + 121 + 494 - 362 - (-411) = 773 \text{ kJ mol}^{-1}$ . However, this approach is more often used to determine the electron affinity of the atom gaining the electron because this quantity is difficult to determine experimentally. Electron affinities for some atoms have been determined only by this procedure rather than experimentally.

If 1 mol of  $\text{Na}^+(g)$  and 1 mol of  $\text{Cl}^-(g)$  are allowed to interact to produce 1 mol of *ion pairs*, the energy released is about  $-439 \text{ kJ}$ . If, on the other hand, 1 mol of  $\text{Na}^+(g)$  and 1 mol of  $\text{Cl}^-(g)$  are allowed to form a mole of *crystal*, the energy released is about  $-773 \text{ kJ}$ . The opposite of this process, the separation of 1 mol of crystal into its gaseous ions, is accompanied by an enthalpy change defined as lattice energy. Note that the ratio of the energy released when the crystal is allowed to form,  $-773 \text{ kJ}$ , to that when ion pairs form,  $-439 \text{ kJ}$ , is about 1.75. This point will be discussed later. As shown previously, it is necessary to know the ionization potential and enthalpy of sublimation for the metal, the dissociation energy and electron affinity of the nonmetal, and the heat of formation of the alkali halide in order to calculate the lattice energy. Table 4.1 shows these values for the alkali metals and the halogens, and the heats of formation of the alkali halides are given in Table 4.2.

If all the interactions in a mole of crystal composed of  $+1$  and  $-1$  ions are taken into account, it is possible to derive the following equation for the lattice energy,  $U$  (in  $\text{kJ/mole}$ ).

$$U = \frac{N_o A e^2}{r} \left( 1 - \frac{1}{n} \right) \quad (4.5)$$

**TABLE 4.1** Thermochemical Data for the Alkali Metals and Halogens

Metal	$I, \text{ kJ mol}^{-1}$	$S, \text{ kJ mol}^{-1}$	Halogen	$E, \text{ kJ mol}^{-1}$	$D, \text{ kJ mol}^{-1}$
Li	518	160	$\text{F}_2$	333	158
Na	494	109	$\text{Cl}_2$	348	242
K	417	90.8	$\text{Br}_2$	324	193
Rb	401	83.3	$\text{I}_2$	295	151
Cs	374	79.9			

$I$  = ionization potential;  $S$  = sublimation enthalpy;  $E$  = electron affinity;  $D$  = dissociation energy.

**TABLE 4.2** Heats of Formation of Alkali Halides, MX

Metal	$-\Delta H_f^\circ, \text{ kJ mol}^{-1}$			
	X = F	X = Cl	X = Br	X = I
Li	605	408	350	272
Na	572	411	360	291
K	563	439	394	330
Rb	556	439	02	338
Cs	550	446	408	351

In this equation,  $N_o$  is Avogadro's number ( $6.022 \times 10^{23} \text{ mol}^{-1}$ ),  $e$  is the charge on the electron ( $4.808 \times 10^{-10} \text{ esu}$ ),  $r$  is the distance between positive and negative ions (in pm), and  $n$  is a constant related to the electron configuration of the ions. The constant  $A$  is known as the *Madelung constant*, and it is determined by how the ions are arranged in the lattice and the number of cations and anions in the formula. Each type of crystal lattice has a different value for this constant. The Madelung constant for NaCl is about 1.75, a value that is equal to the ratio of the energy when a mole of crystal forms to that when a mole of ion pairs forms. Essentially, the Madelung constant is a factor that takes into account the fact that in a lattice each cation interacts with more than one anion and vice versa. Thus, there is additional energy released when a lattice is formed compared to that if only ions pairs form. Table 4.3 shows the Madelung constants for some common crystal types.

As ions get closer together, there are some forces of repulsion even though the ions may be oppositely charged. This repulsion arises from the interaction of the electron clouds in the separate ions, and it is usually expressed, as was done by Born, in terms of a factor involving  $1/r^n$ . Thus, the repulsion term is usually written as  $B/r^n$  where  $B$  is a proportionality constant and  $n$  has the value 5, 7, 9, 10, or 12 depending upon whether the ion has the electron configuration of helium, neon, argon, krypton, or xenon, respectively. The attraction between a mole of  $+1$  ions and a mole of  $-1$  ions is  $-N_o A e^2 / r$ , so the addition of the repulsion gives the expression for the lattice energy shown in Eq. (4.6).

$$U = \frac{-N_o A e^2}{r} + \frac{B}{r^n} \quad (4.6)$$

When an equilibrium distance is reached, the forces of attraction and repulsion give rise to the most favorable or minimum energy. At this point, the energy versus distance curve goes through a minimum. Because of this,  $B$  can be eliminated by differentiating Eq. (4.6) with respect to  $r$  and setting the derivative equal to zero,

$$\frac{\partial U}{\partial r} = 0 = \frac{N_o A e^2}{r^2} - \frac{nB}{r^{n+1}} \quad (4.7)$$

**TABLE 4.3** Madelung Constants for Crystal Lattices

Crystal Type	Madelung Constant <sup>a</sup>
Sodium chloride	1.74756
Cesium chloride	1.76267
Zinc blende	1.63806
Wurtzite	1.64132
Rutile	2.408
Fluorite	2.51939

<sup>a</sup>Does not include the factor of  $Z_+Z_-$  for ions having charges other than  $+1$  and  $-1$ . Does not include the fact that rutile and fluorite structures have twice as many anions as cations.

so that

$$B = \frac{N_0 A e^2 r^{n-1}}{n} \quad (4.8)$$

Substituting this value for  $B$  in Eq. (4.6) and changing signs because  $U$  is defined as the energy required to *separate* a mole of the crystal into gaseous ions (so it has a positive sign) yields the Born–Mayer equation,

$$U = \frac{N_0 A e^2}{r} \left( 1 - \frac{1}{n} \right) \quad (4.9)$$

This equation is identical to Eq. (4.5) given earlier for calculating the lattice energy. When this equation is written to describe a crystal for ions that have charges of  $Z_a$  and  $Z_c$  rather than  $+1$  and  $-1$ , it becomes

$$U = \frac{N_0 A Z_a Z_c e^2}{r} \left( 1 - \frac{1}{n} \right) \quad (4.10)$$

Figure 4.1 shows the calculated lattice energy as a function of the distance between  $+1$  and  $-1$  ions for crystals having the NaCl-type lattice when  $n = 7$ .

It is possible to obtain other semiempirical expressions that express the lattice energy in terms of ionic radii, charges on the ions, etc. One of the most successful of these is the *Kapustinskii equation*,

$$U(\text{kJ mol}^{-1}) = \frac{120,200 m Z_a Z_c}{r_c + r_a} \left( 1 - \frac{34.5}{r_c + r_a} \right) \quad (4.11)$$

where  $r_c$  and  $r_a$  are ionic radii in pm, and  $m$  is the number of ions in the formula for the compound (2 for NaCl, 3 for  $\text{MgCl}_2$ , etc.). This equation works very well for substances that are essentially ionic (e.g., LiF, NaCl, etc.). The agreement between the lattice energy calculated using Eq. (4.11) and that obtained from thermochemical data is rather poor for crystals such as AgI where the charges on the ions are distributed over relatively large volumes and van der Waals forces become more important. Equation (4.11) is based on a purely electrostatic model and does not give good agreement when the ions are highly polarizable and the crystal is somewhat covalent. This will be discussed in more detail later.

### 4.1.2 Radius Ratio Effects

The relative size of the ions forming a crystal lattice has a great influence upon what type of lattice forms. It is necessary for ions to “touch” ions having the opposite charge in order for there to be a net attraction and the arrangement of ions in the crystal to be a stable one. Otherwise, the ions touch only ions of the same charge and repulsion causes this arrangement to be unstable. Ions are considered to be hard spheres in this approach. If the two ions are about equal in size, it may be possible for six or more ions of opposite charge to surround a given ion. Consider the arrangements shown in Figure 4.2.

The arrangement of ions shown in (a) will be stable whereas that shown in (b) will not. In arrangement (b), only the negative ions touch each other and give rise to repulsion that will cause the arrangement to be unstable. If the negative ions are large enough that six of them cannot surround a given positive ion, it may be possible for four of the anions to all touch the positive ion and lead to a stable arrangement. It is apparent that the critical factor is the radius ratio,  $r_c/r_a$ , that expresses

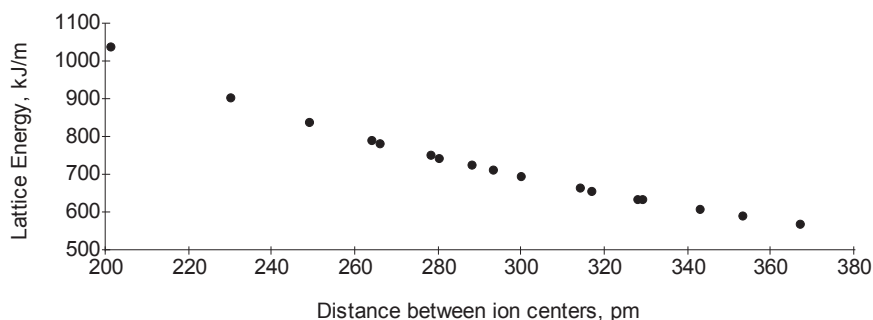
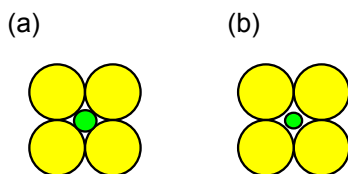


FIGURE 4.1 The variation in lattice energy as a function of the internuclear distance for a sodium chloride-type lattice.



**FIGURE 4.2** Cations in octahedral holes surrounded by anions. In (a), the anions all contact the cation but in (b), the anions contact only each other.

the relative size of the ions. In order to use this approach, the radii of the ions must be known, and [Table 4.4](#) shows radii of many common ions. For polyatomic ions, the radii are based on calculations using thermochemical cycles rather than on X-ray analysis of crystals so they are referred to as *thermochemical radii*. These values are very useful for determining lattice energies using the Kapustinskii equation, the use of which requires values for the ionic radii but not the Madelung constant.

Consider a positive ion of radius  $r_c$  surrounded by six anions of radius  $r_a$  in an octahedral arrangement as shown in [Figure 4.3](#). An anion lies on either side of the plane of the page. With this arrangement, the four anions shown just touch the cation and just touch each other. From the geometric arrangement, it can be seen that  $\theta = 45^\circ$  and  $S = r_c + r_a$ . It follows directly from [Figure 4.3](#) that

$$\cos 45^\circ = \frac{\sqrt{2}}{2} = \frac{r_a}{S} = \frac{r_a}{r_c + r_a} \quad (4.12)$$

TABLE 4.4 Ionic Radii					
Singly Charged		Doubly Charged		Triply Charged	
Ion	$r$ , pm	Ion	$r$ , pm	Ion	$r$ , pm
$\text{Li}^+$	60	$\text{Be}^{2+}$	30	$\text{Al}^{3+}$	50
$\text{Na}^+$	98	$\text{Mg}^{2+}$	65	$\text{Sc}^{3+}$	81
$\text{K}^+$	133	$\text{Ca}^{2+}$	94	$\text{Ti}^{3+}$	69
$\text{Rb}^+$	148	$\text{Sr}^{2+}$	110	$\text{V}^{3+}$	66
$\text{Cs}^+$	169	$\text{Ba}^{2+}$	129	$\text{Cr}^{3+}$	64
$\text{Cu}^+$	96	$\text{Mn}^{2+}$	80	$\text{Mn}^{3+}$	62
$\text{Ag}^+$	126	$\text{Fe}^{2+}$	75	$\text{Fe}^{3+}$	62
$\text{NH}_4^+$	148	$\text{Co}^{2+}$	72	$\text{N}^{3-}$	171
$\text{F}^-$	136	$\text{Ni}^{2+}$	70	$\text{P}^{3-}$	212
$\text{Cl}^-$	181	$\text{Zn}^{2+}$	74	$\text{As}^{3-}$	222
$\text{Br}^-$	195	$\text{O}^{2-}$	145	$\text{Sb}^{3-}$	245
$\text{I}^-$	216	$\text{S}^{2-}$	190	$\text{PO}_4^{3-}$	238
$\text{H}^-$	208	$\text{Se}^{2-}$	202	$\text{SbO}_4^{3-}$	260
$\text{ClO}_4^-$	236	$\text{Te}^{2-}$	222	$\text{BiO}_4^{3-}$	268
$\text{BF}_4^-$	228	$\text{SO}_4^{2-}$	230		
$\text{IO}_4^-$	249	$\text{CrO}_4^{2-}$	240		
$\text{MnO}_4^-$	240	$\text{BeF}_4^{2-}$	245		
$\text{NO}_3^-$	189	$\text{CO}_3^{2-}$	185		
$\text{CN}^-$	182				
$\text{SCN}^-$	195				

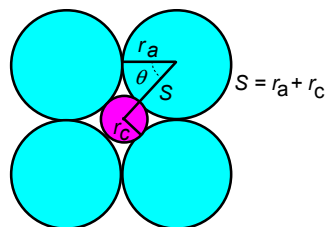


FIGURE 4.3 A cation in an octahedral hole of anions.

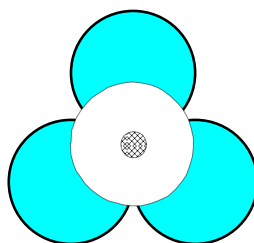


FIGURE 4.4 A cation (cross-hatched circle) in a tetrahedral hole.

from which we obtain the value  $r_c/r_a = 0.414$ . Thus, if the cation is smaller than  $0.414 r_a$ , this arrangement is not likely to be stable because the cation will fit in the octahedral hole without touching the anions. Therefore, the radius ratio of 0.414 represents the lower limit of size where an octahedral arrangement of anions around the cation is expected. The cation might, however, fit in a tetrahedral hole as shown in Figure 4.4.

The type of analysis shown for the tetrahedral case leads to the conclusion that the cation cannot touch all the anions unless  $r_c \geq 0.225 r_a$ . As a result, when  $r_c/r_a$  is in the range 0.225–0.414 a tetrahedral arrangement of anions around the cation is expected. Table 4.5 summarizes the results of similar calculations for other arrangements of ions.

Using the ionic radii shown in Table 4.4, it is found that nine of the alkali halides should have structures other than the NaCl type. However, only three of these compounds (CsCl, CsBr, and CsI) have a different structure, that being the CsCl type. Apparently, factors other than the radius ratio affect the crystal structure of the alkali halides. That this is so is not especially surprising. One factor responsible is that the ionic radius itself is somewhat dependent on the coordination number of the ion. Another factor that has been ignored is the polarizability of the ions. The ions have been considered as hard spheres when in fact, especially for the larger ions, they are somewhat polarizable and as a result of the charge separation induced, they can have an appreciable covalency to the bonding. The effects of the additional van der Waals forces and covalency can be seen when the lattice energies of the silver halides shown in Table 4.6 are considered.

TABLE 4.5 Stable Arrangements of Ions Related to Radius Ratio Values

$r_c/r_a$	Cation Environment	Number of Nearest Neighbors	Example
$\geq 0.732$	Cubic	8	CsCl
0.414–0.732	Octahedral	6	NaCl
0.225–0.414	Tetrahedral	4	ZnS
0.155–0.225	Trigonal	3	—
$< 0.155$	Linear	2	—

**TABLE 4.6** Lattice Energies of Silver Halides

Compound	Lattice Energy, $\text{kJ mol}^{-1}$	
	Calculated (Eq. (4.11))	Experimental
AgF	816	912
AgCl	715	858
AgBr	690	845
AgI	653	833

Of course, the examples shown in Table 4.6 are those where a large covalent contribution is expected. However, these data show conclusively that a hard sphere ionic model does not account for all the properties of some materials even though they are *predominantly* ionic.

The value for  $r_c/r_a$  for CsCl is 0.934 and as expected the structure has eight  $\text{Cl}^-$  ions surrounding each  $\text{Cs}^+$  ion. However, it is interesting to note that even CsCl has the sodium chloride structure at temperatures above  $445^\circ\text{C}$ . Some of the other alkali halides that normally have the sodium chloride structure exhibit the CsCl structure when subjected to very high pressure.

Basically, ionic radii are determined by X-ray studies on crystals. Such studies serve only to determine the distances between ion centers, not the radii of the individual ions. One independent determination is required and when it is available, the others can be obtained from measured interionic distances. Usually, the values for ionic radii show a relative constancy. That is, the difference between the radius of  $\text{Cl}^-$  and  $\text{Br}^-$  is usually very nearly the same for a series of similar ionic metal halides. Because there is some covalent contribution to polar covalent bonds, establishing the exact boundary and hence the radius of an ion is simple in principle but more complicated in practice. As a result, a rather wide range of values for the radii exists for some ions, depending on the premises on which the determinations are made. It is also known that the radius of an ion depends somewhat on the ligancy (number of nearest neighbors or coordination number). However, the ionic radius does not vary greatly with change in coordination number with the variation usually being only about 5% at most.

Although the predictions from radius ratios are not completely reliable, they are correct in most cases. The deviations usually occur when the radius ratio falls near one of the limits. For example, if  $r_c/r_a$  is about 0.405, it might be expected that the environment around the cation would be tetrahedral. A very slight change in overall energy might cause the actual structure to have an octahedral environment around the cation. If  $r_c/r_a$  is 0.58, the environment around the cation is virtually always octahedral because that value is near the middle of the 0.414–0.732 range.

### 4.1.3 Crystal Structures

A relatively small number of crystal structures describe a vast number of ionic compounds. In these structures, the environment around the cation and anion are described in terms of the number of nearest neighbors (also known as the coordination number). The coordination numbers for the anion and cation for some common crystal structures of binary inorganic compounds and some examples of compounds that have that crystal type are shown in Table 4.7. The stoichiometry of the compound may dictate that there will be different coordination numbers for the anion and cation. For example, the structure of  $\text{CaF}_2$  is like that of CsCl but with every other cation site being vacant because there are two anions per cation.

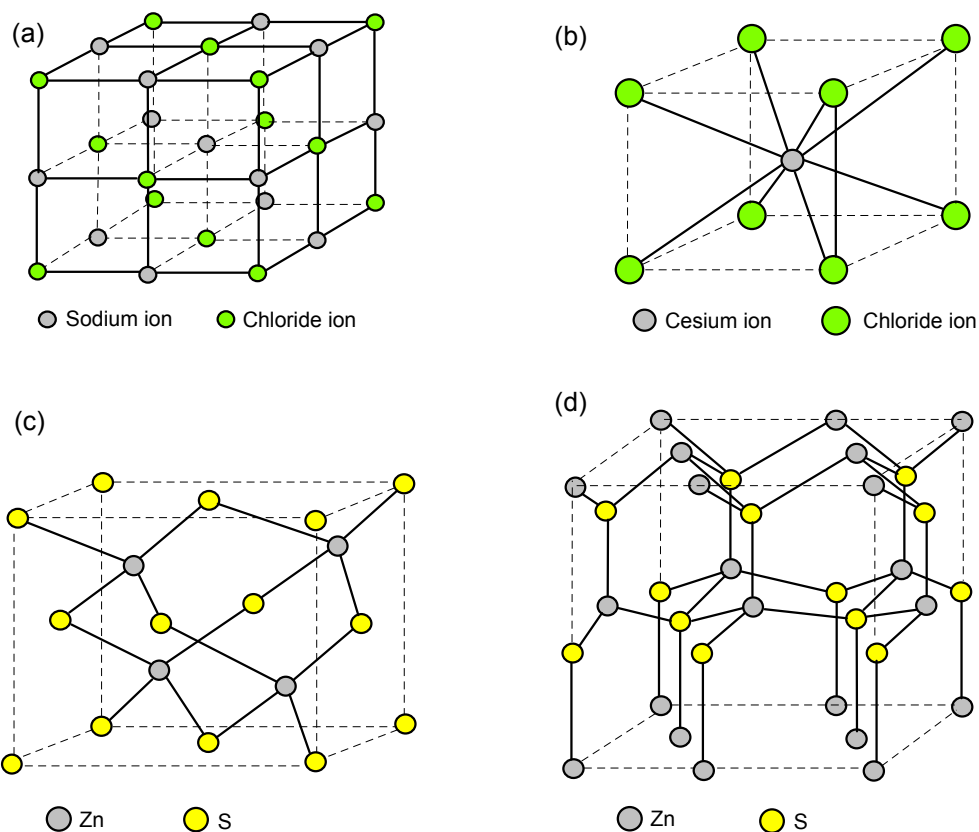
When describing the structures of inorganic compounds, it is important to have in mind the spatial arrangement of the ions. The structural types of some of the common binary inorganic compounds are shown in Figures 4.5 and 4.6. From the number of compounds that exhibit the various crystal types, it is readily apparent that these are very important structural *types* for inorganic materials. They are also the arrangements for compounds that contain cations or anions composed of more than one atom. For example,  $\text{KNO}_3$  and  $\text{CaCO}_3$  have the NaCl-type structure.

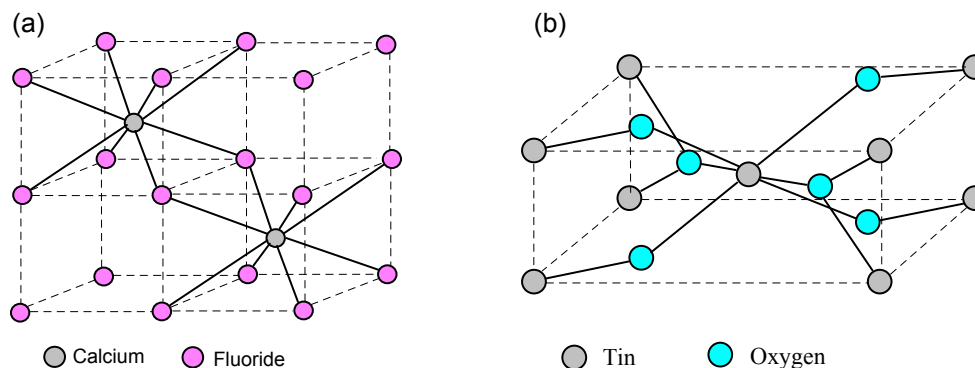
Another important structural type is that of  $\text{ReO}_3$ . Because there are three anions per cation, its structure must be different from any of those shown thus far because none of those previously shown have that ratio. The  $\text{ReO}_3$  structure is shown in Figure 4.7.



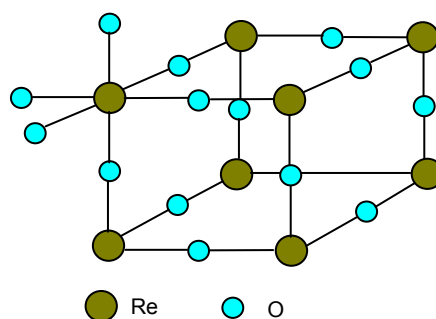
**TABLE 4.7** Structural Information and Examples of Some Common Crystal Types

Crystal Type	Cation Coordination Number	Anion Coordination Number	Examples
NaCl	6	6	NaCl, NaH, CsF, AgCl, NH <sub>4</sub> I, MgO, MgS, TiN, CaCO <sub>3</sub>
CsCl	8	8	CsCl, CsBr, CsI, TlCl, NH <sub>4</sub> Cl, NH <sub>4</sub> Br, K <sup>+</sup> SbF <sub>6</sub> <sup>-</sup>
Zinc blende	4	4	ZnS, BeS, CdS
Wurtzite	4	4	ZnS, MnS, CdS
Fluorite	8	4	CaF <sub>2</sub> , SrF <sub>2</sub> , BaF <sub>2</sub>
Antifluorite	4	8	Li <sub>2</sub> S, Na <sub>2</sub> S, K <sub>2</sub> S
Rutile	6	3	TiO <sub>2</sub> (rutile), SnO <sub>2</sub> , PbO <sub>2</sub> , MnO <sub>2</sub> , MgF <sub>2</sub> , ZnF <sub>2</sub> , CoF <sub>2</sub> , MnF <sub>2</sub> , FeF <sub>2</sub> , NiF <sub>2</sub>
ReO <sub>3</sub>	6	2	ReO <sub>3</sub> , VF <sub>3</sub> , WO <sub>3</sub> , RuF <sub>3</sub>
Perovskite	—	—	CaTiO <sub>3</sub> , KMgF <sub>3</sub> , RbIO <sub>3</sub>

**FIGURE 4.5** Structures for some common crystal types in which the ratio of cation to anion is 1:1. (a) is the NaCl or rock salt structure, (b) is the CsCl structure, (c) is the zinc blende structure for ZnS, and (d) is the wurtzite structure for ZnS.



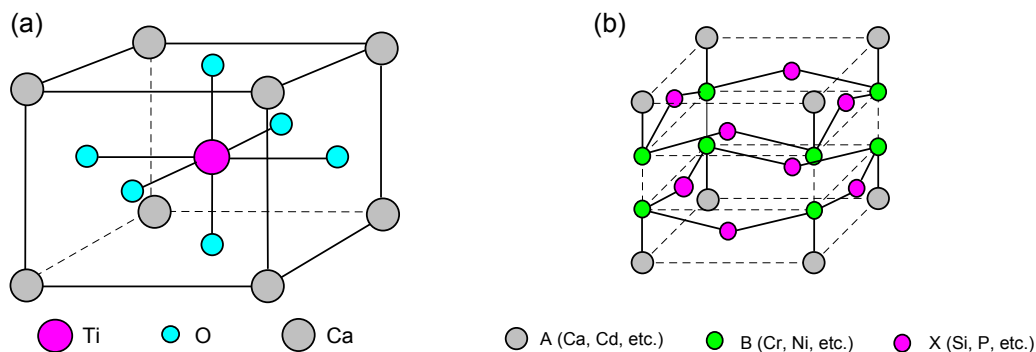
**FIGURE 4.6** Structures for some common crystal types in which the ratio of cation to anion is 1:2. (a) is the fluorite structure. The fluorite structure is a common structural type for 1:2 compounds. If the compound has a 2:1 formula, the role of the cation and anion are reversed, and this gives the antifluorite structure that is shown by compounds such as  $\text{Na}_2\text{S}$ . (b) is the  $\text{TiO}_2$  or rutile structure.



**FIGURE 4.7** The  $\text{ReO}_3$  structural type in which the ratio of cations to anions is 1:3. Note the octahedral arrangement of anions around each cation.

As has been mentioned, ternary compounds having polyatomic ions such as  $\text{NO}_3^-$ ,  $\text{CO}_3^{2-}$ ,  $\text{NH}_4^+$ , etc., often have the same types of structures as binary compounds in which a polyatomic ion occupies a lattice site as a unit. The mineral *perovskite*,  $\text{CaTiO}_3$ , calcium titanate, however, is a somewhat different type of ternary compound that has the structure shown in [Figure 4.8\(a\)](#).

Most ternary compounds are oxides, and the general formula  $\text{ABO}_3$  corresponds to many compounds because  $\text{A} = \text{Ca}$ ,  $\text{Sr}$ ,  $\text{Ba}$ , etc., and  $\text{B} = \text{Ti}$ ,  $\text{Zr}$ ,  $\text{Al}$ ,  $\text{Fe}$ ,  $\text{Cr}$ ,  $\text{Hf}$ ,  $\text{Sn}$ ,  $\text{Cl}$ , or  $\text{I}$ . Accordingly, *perovskite* is an important structural type, but it is not always easy to visualize the environment around  $\text{Ca}$  in the structure. Fortunately, there is a procedure for describing the bonding in crystals that enables us to determine what the arrangement around  $\text{Ca}$  must be. We will now describe this procedure.



**FIGURE 4.8** Structures for some common ternary compounds. (a) is the perovskite structure for which the general formula is  $\text{ABX}_3$ . The  $\text{Ti}$  is surrounded octahedrally by six oxide ions; each  $\text{O}$  is surrounded by two  $\text{Ti}$  ions and four  $\text{Ca}$  ions. Each  $\text{Ca}$  is surrounded by 12  $\text{O}$  ions. (b) is the structure of compounds that have the formula  $\text{AB}_2\text{X}_2$ .

In the structure of NaCl, there are six  $\text{Cl}^-$  ions surrounding each  $\text{Na}^+$  ion. Because the valence of Na is 1, each of the six  $\text{Cl}^-$ – $\text{Na}^+$  interactions must be  $1/6$  of a bond so that the total adds up to the valence of  $\text{Na}^+$ . Likewise, if each interaction between  $\text{Cl}^-$  and  $\text{Na}^+$  is  $1/6$  of a bond, there must be six  $\text{Na}^+$  around each  $\text{Cl}^-$  to add up to the valence of  $\text{Cl}^-$  which is also 1 (without regard to sign). The principle here is that *the sum of the interactions from the nearest neighbors of an ion must add up to the valence of the ion*. The electrostatic bond character is the fraction of a bond that the interaction represents. This enables some of the features of the structure to be deduced as will now be illustrated.

Consider the  $\text{CaF}_2$  crystal. Whatever the bond character between  $\text{Ca}^{2+}$  and  $\text{F}^-$  is, there must be twice as many bonds to  $\text{Ca}^{2+}$  as a result of it having twice the magnitude of the valence of  $\text{F}^-$ . The  $\text{TiO}_2$  crystal represents the same situation. There must be twice as many bonds to  $\text{Ti}^{4+}$  as there are to  $\text{O}^{2-}$  because of the valences of the species. Examination of the structures of  $\text{CaF}_2$  and  $\text{TiO}_2$  shown in Figure 4.6 shows that this is the case.

The unit cell of perovskite,  $\text{CaTiO}_3$ , is shown in Figure 4.8(a). The  $\text{Ti}^{4+}$  ion resides in the center of the cube, the  $\text{O}^{2-}$  ions are shared on the faces, and the  $\text{Ca}^{2+}$  ions are located on the corners. Note that there are six  $\text{O}^{2-}$  ions surrounding the  $\text{Ti}^{4+}$  ion. Because it takes six bonds to  $\text{O}^{2-}$  ions to satisfy the +4 valence of  $\text{Ti}^{4+}$ , each bond must be  $4/6$  or  $2/3$  in electrostatic bond character. Also, each oxide ion is located on a face of the cube between two  $\text{Ti}^{4+}$  ions that are located in the centers of the cubes joined at that face. Thus, two  $\text{Ti}^{4+}$ – $\text{O}^{2-}$  bonds must represent  $2(2/3) = 4/3$  of a bond to each oxide ion. The remainder of the valence of the  $\text{O}^{2-}$  ( $2 - 4/3 = 2/3$ ) is made up by the  $\text{Ca}^{2+}$  ions on the corners of the face of the cube where the oxide ion resides. It is readily apparent that there are four  $\text{Ca}^{2+}$  nearest neighbors for each  $\text{O}^{2-}$ . However, it is not so obvious how many  $\text{O}^{2-}$  nearest neighbors each  $\text{Ca}^{2+}$  has, but it is quite easy to determine the number. Four  $\text{Ca}^{2+}$ – $\text{O}^{2-}$  interactions add up to  $2/3$  of a bond so each of these interactions must have an electrostatic bond character such that four of them equal  $2/3$  of a bond. From this it can be seen that  $4x = 2/3$  so that  $x = 1/6$ . Because the valence of  $\text{Ca}^{2+}$  is 2, it would take 12  $\text{O}^{2-}$  ions surrounding the  $\text{Ca}^{2+}$  ions, each with a bond character of  $1/6$ , to satisfy the valence of +2. So, in the perovskite structure, the  $\text{Ti}^{4+}$  has a coordination number of 6,  $\text{O}^{2-}$  has a coordination number of 6 (four  $\text{Ca}^{2+}$  and two  $\text{Ti}^{4+}$ ), and  $\text{Ca}^{2+}$  has a coordination number of 12. The perovskite structure shown in Figure 4.8(a) verifies these predictions. This procedure is of great utility in describing the environment of ions in a crystal, and we will make use of it in several cases throughout this book.

There are several hundred possible compounds having the formula  $\text{AB}_2\text{X}_2$  where A is a +2 metal (Ca, Ba, Cd, etc.), B is a transition metal (Ni, Co, Zr, Fe, etc.), and X is a main group element such as N, P, or Si. These compounds have the general structure shown in Figure 4.8(b). In compounds that have the general formula  $\text{AB}_2\text{X}_2$ , each main group atom resides at the apex of a square base pyramid of transition metal atoms. The distance between the two layers of X atoms varies considerably depending on the nature of B in the formula  $\text{AB}_2\text{X}_2$ . It has been interpreted that this variation is related to the ease of oxidizability of the metal B. For example, if we consider the compounds  $\text{SrCu}_2\text{P}_2$  and  $\text{SrFe}_2\text{P}_2$ , the P–P distances are 229 and 343 pm, respectively. On the basis of the bond character procedure described earlier, increasing the P–B bond character would of necessity decrease the P–P bond character because the number of nearest neighbors and valences has not changed. With a fixed total valence of P, increasing the bond character between the P and B atoms would leave less of the total valence on P to be made up by the P–P bonds. In fact, the distance of 343 pm between P atoms in  $\text{SrFe}_2\text{P}_2$  has been interpreted as meaning that there is no P–P bonding in that compound.

We have shown some of the structures exhibited by inorganic compounds. Others will be described as we examine the chemistry of particular elements and their compounds. Also, there are a great many covalent extended structures (e.g., diamond, graphite, quartz, boron nitride, etc.) that will be discussed in later chapters as we deal with the chemistry of those materials. The discussion presented here is sufficient to serve as a basis for describing the behavior and properties of many inorganic substances. This is an important objective for understanding descriptive chemistry.

## 4.2 INTERMOLECULAR INTERACTIONS

Although Chapter 3 dealt with covalent bonds between atoms and up to this point this chapter has been devoted to ionic bonding, there are other forces that are important in inorganic chemistry. Some description will now be given of the types of forces that exist between molecules because these forces determine many of the physical and chemical properties of the substances. It is clear that some force must hold molecules together in dry ice, liquid water, and all other *molecular* solids and liquids or else they would not exist in condensed phases. Generally, such intermolecular forces amount to only perhaps  $2\text{--}20 \text{ kJ mol}^{-1}$  so that they are much weaker than the usual valence forces. Although they are weak, these forces have a great influence on the properties of materials so some familiarity with their nature is essential. In a general way, the boiling points of liquids reflect the strengths of intermolecular forces because the process of boiling separates molecules from each other. There are several types of intermolecular forces, and the first three types of these forces are usually referred to collectively as *van der Waals forces*.

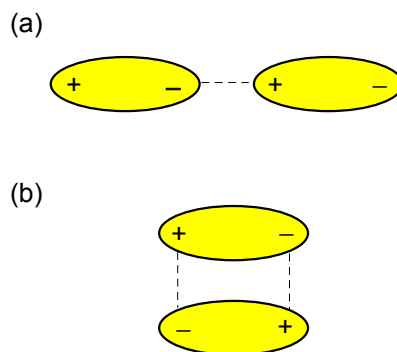
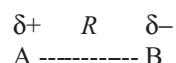


FIGURE 4.9 Interaction of dipoles by the (a) parallel and (b) antiparallel modes.

### 4.2.1 Dipole–Dipole Forces

Covalent bonds can have appreciable polarity due to the unequal sharing of electrons by atoms that have different electronegativities. For most types of bonds, this charge separation amounts to only a small percentage of an electron charge. For example, in HI it is about 5%, but in HF where the difference in electronegativity is about 1.8 units, it is about 44%.

In order to show how dipole–dipole forces arise, let us consider a polar molecule that can be represented as



where  $\delta+$  and  $\delta-$  represent the fraction of an electronic charge residing on the positive and negative ends, respectively. When polar molecules are allowed to approach each other, there will be an electrostatic interaction between them. The actual energy of the interaction will depend on the orientation of the dipoles with respect to each other. Two limiting cases can be visualized as shown in Figure 4.9.

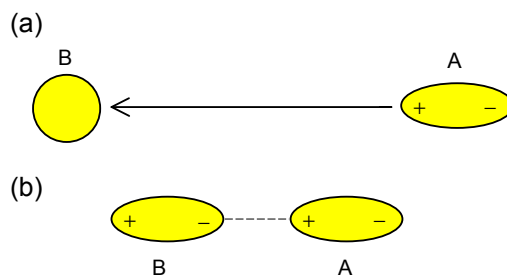
By assuming an averaging of all possible orientations, the energy of interaction,  $E_D$  can be shown to be

$$E_D = -\frac{2\mu^4}{3kTR^6} \quad (4.13)$$

where  $\mu$  is the dipole moment,  $R$  is the average distance of separation,  $k$  is Boltzmann's constant, and  $T$  is the temperature (K). On the basis of this interaction, it is expected that polar molecules should associate to some extent, either in the vapor state or in solvents of low dielectric constant. Dipole association in a solvent having a low dielectric constant leads to an abnormal relationship between the dielectric constant of the solution and the concentration of the polar species. Although the procedure will not be shown, it is possible to calculate the association constants for such systems from the dielectric constants of the solutions. If the solvent has a high dielectric constant and is polar, it may solvate the polar solute dipoles thus preventing association which forms aggregates. Consequently, the association constants for polar species in solution are always dependent on the solvent used.

### 4.2.2 Dipole-Induced Dipole Forces

A molecule that has a dipole moment of zero in the absence of any interactions with other molecules can have a charge separation induced in it by interaction with a polar molecule in which there is charge separation. This behavior is illustrated in Figure 4.10. The interaction between molecules A and B is now energetically favorable as a result of the dipole moment induced in B. The susceptibility of B to this type of induction depends upon the extent to which its electronic cloud can be distorted. A measure of this distortion is given by the *electronic polarizability*,  $\alpha$ . In a general way, the polarizability is related to the size of the molecule and the number of electrons it contains. For example, the polarizability of CO is about  $2.0 \text{ \AA}^3/\text{molecule}$  whereas that of  $\text{CCl}_4$  is about  $10.5 \text{ \AA}^3/\text{molecule}$ . However, the manner in which the electrons are distributed is also important because the electrons may be easier to move in one molecule than in another. Benzene and  $\text{CCl}_4$  had approximately equal polarizabilities (and boiling points) even though the molecular weight of  $\text{CCl}_4$  is approximately twice that of benzene. This is the result of the  $\pi$  electron cloud in benzene being mobile and polarizable whereas the electrons (both bonding and unshared pairs) in  $\text{CCl}_4$  are more localized.



**FIGURE 4.10** In (a), a polar molecule approaches a nonpolar molecule. In (b), charge separation is induced in the nonpolar species by the nearby polar molecule.

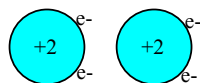
The energy of interaction between a molecule having a dipole moment of  $\mu$  and another having a polarizability of  $\alpha$  is given as

$$E_I = -\frac{2\mu\alpha^2}{R^6} \quad (4.14)$$

This type of interaction can also occur between a charged ion and a nonpolar molecule.

### 4.2.3 London Dispersion Forces

Dipole–dipole and dipole-induced dipole interactions result from the attraction of species having charge separations. However, molecules having no charge separation can be liquefied. Helium,  $\text{H}_2$ ,  $\text{N}_2$ ,  $\text{O}_2$ , and other nonpolar molecules still interact weakly. The nature of this interaction is determined by the fact that even for nonpolar atoms and molecules the electrons are not always distributed symmetrically. It is certainly possible that for He the two electrons can be found at some particular instant on the same side of the nucleus.



There is an *instantaneous* dipole that will cause there to be an instantaneous change in the electron distribution in the neighboring atom. As a result, there is a weak attraction between the nucleus in one molecule and the electrons in another. The ability to have the electrons shifted in this way is related to the polarizability,  $\alpha$ , of the molecule. This type of force is called the *London* or *dispersion* force, and the energy of this type of interaction,  $E_L$ , can be expressed by the equation

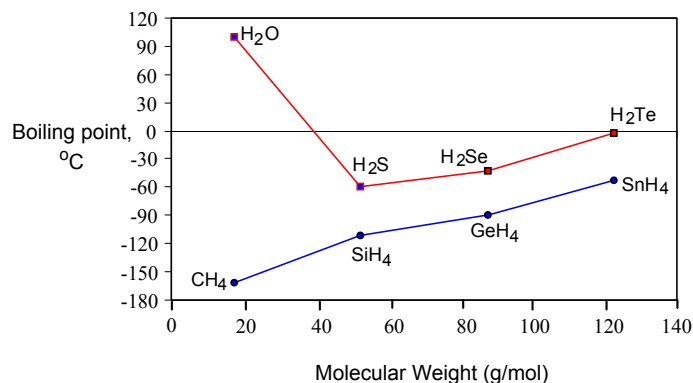
$$E_L = -\frac{3I\alpha^2}{4R^6} \quad (4.15)$$

where  $R$  is the distance of separation,  $\alpha$  is polarizability, and  $I$  is the ionization potential. For many molecules, the ionization potential is essentially constant. For example,  $\text{SO}_2$ ,  $\text{CH}_3\text{OH}$ , and benzene have the ionization potentials of 11.7, 10.85, and 9.24 eV/molecule (1130, 1050, and 892  $\text{kJ mol}^{-1}$ ), respectively. For a wide range of molecules, the range of ionization potentials is about 850–1200  $\text{kJ mol}^{-1}$ .

Because the London interaction energy depends on the polarizability, it increases in a general way with molecular size and number of electrons. Delocalized electrons, as in benzene, cause the polarizability, and hence London forces, to be large. As a result,  $\text{C}_6\text{H}_6$ , with a molecular mass of 78, has a boiling point as high as that of  $\text{CCl}_4$  that has a molecular mass of 154. However, for many nonpolar materials, the boiling point generally increases as the molecular mass increases.

### 4.2.4 Hydrogen Bonding

Because of its low electronegativity (about 2.1), bonds between hydrogen and atoms such as F, O, N, Cl, or S have considerable polarity. Therefore, the hydrogen atom in one molecule can be attracted to a pair of electrons in another. Hydrogen is unique in that when one covalent bond is formed, the nucleus has no other electrons to shield it. The electron donor atom may be in the same molecule giving rise to *intramolecular* hydrogen bonds, or in another molecule giving rise to *intermolecular* hydrogen bonds. Although at first glance it might appear to be a rather special and limited type of



**FIGURE 4.11** The boiling points of the hydrogen compounds of the Group IVA and Group VIA elements as a function of molecular weight. Note the effect of the strong hydrogen bonding in the case of H<sub>2</sub>O.

interaction, it is in fact quite widely occurring. The number of molecules containing O–H or N–H bonds is very large, and the number of molecules that contain atoms capable of functioning as electron pair donors is also large. As a result, hydrogen bonds occur in many systems, including natural and biologically occurring materials such as starch, cellulose, DNA, proteins, leather, etc.

Thousands of research papers and several books are available on the subject of hydrogen bonding. This large body of literature serves to illustrate the fundamental nature of hydrogen bonding and the importance of understanding the effects on properties that hydrogen bonding produces. Hydrogen bonds may form between molecules in a pure liquid (e.g., water, alcohols, or HF) or between a donor and acceptor when both are dissolved in an inert solvent. There is extensive hydrogen bonding in liquid HF and HCN where these materials are essentially polymeric as a result of hydrogen bonding.



There is also an interaction between the OH group of an alcohol and the  $\pi$ -electrons in molecules such as benzene when both are dissolved in an inert solvent.

Hydrogen bonding also occurs in a great number of solids. For example, NH<sub>4</sub>Cl and other ammonium compounds have N–H $\cdots$ X hydrogen bonds where X is an anion functioning as an electron pair donor. In some cases, breaking part or all of the hydrogen bonds results in a change in crystal structure (a phase transition).

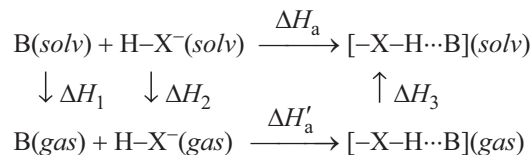
Hydrogen bonding affects the physical properties of liquids. The association of molecules causes them to be more strongly attracted to each other than would be expected on the basis of increasing molecular weight. As shown in Figure 4.11 this is readily apparent when the hydrogen compounds of Group VIA and IVA are considered.

Intermolecular hydrogen bonding represents an interaction by van der Waals forces that leads to association of molecules. The number of experimental methods for studying the effects of hydrogen bonding is quite large and includes all the classical methods for studying associated liquids. Melting and boiling points, dielectric behavior, vapor pressure, thermal conductivity, index of refraction, viscosity, and solubility are all used to investigate properties of hydrogen-bonded substances. Two of the most important experimental methods are infrared spectroscopy and nuclear magnetic resonance. Some of the important characteristics and consequences of hydrogen bonding are discussed in the following sections.

1. *Hydrogen bond energies.* Hydrogen bonds are frequently classified in terms of the enthalpy of formation of the bonds. Table 4.8 shows representative values for three arbitrary categories that are used to classify hydrogen bonds.

Classification	$-\Delta H$ , kJ mol <sup>-1</sup>	Example
Weak	<12	2-chlorocyclohexanol
Normal	12–40	ROH $\cdots$ NC <sub>5</sub> H <sub>5</sub> , etc.
Strong	>40	[F $\cdots$ H $\cdots$ F] <sup>-</sup>

Part of the difficulty in assessing hydrogen bond enthalpies is in determining the role of the solvent. If the solvent interacts strongly with the two molecules forming the hydrogen bond, the interaction between solute molecules will be hindered. In such cases, solvation competes with the formation of the hydrogen bond. The energies of interactions between solute molecules and the solvation of the species can be seen from the following cycle.

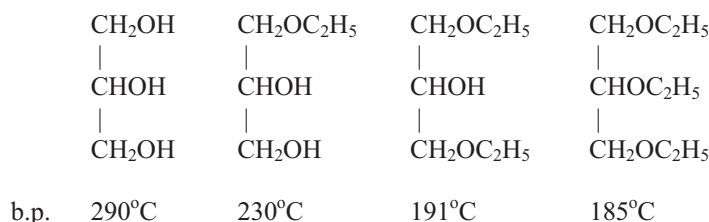


In this cycle, B represents the electron pair donor molecule, X-H represents the species containing the polar bond to hydrogen, and  $[-\text{X-H}\cdots\text{B}]$  is the complex (or adduct) formed between them. Bond enthalpies normally refer to strengths of bonds in gaseous molecules. Therefore, the enthalpy we want is  $\Delta H'_a$ . However, this can seldom be measured directly because most hydrogen-bonded adducts are not sufficiently stable. For example,  $\text{CH}_3\text{OH}$  bonds to pyridine,  $\text{C}_5\text{H}_5\text{N}$ , but the complex is too unstable to exist at  $116^\circ\text{C}$ , the boiling point of pyridine. Accordingly, such interactions are studied in "inert" solvents such as heptane. Carbon tetrachloride with its unshared pairs of electrons on the chlorine atoms is not totally inert. It is apparent that the enthalpy of adduct formation,  $\Delta H_a$ , will be solvent dependent unless  $|\Delta H_3| = |\Delta H_1 + \Delta H_2|$ , where  $\Delta H_1$  and  $\Delta H_2$  are the heats of solvation of the electron donor and the hydrogen compound, respectively, and  $\Delta H_3$  is the heat of solvation of the complex. The enthalpy of the formation of the adduct in the gas phase, which would give the actual strength of the hydrogen bond, will be different from that in solution unless the solvation energies cancel. In many cases, this is questionable because of the electron donor properties of the solvent. There is no doubt that many of the hydrogen bond enthalpies reported in the literature may be in error because of this. Moreover, it has been shown that the extent of self-association of compounds such as alcohols is strongly affected by the solvent used. Molecular association in general depends on the nature of the medium that solvates the species and serves to separate them.

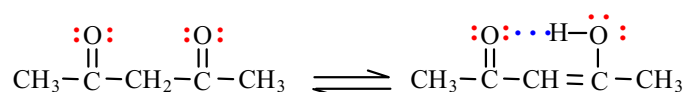
2. *Effects of hydrogen bonding on physical and chemical properties.* Hydrogen bonding produces many physical and chemical effects. The added intermolecular interaction often produces a drastic effect on melting and boiling points. For example,  $\text{H}_2\text{O}$  boils at  $100^\circ\text{C}$  and  $\text{H}_2\text{S}$  boils at  $-61^\circ\text{C}$ .  $\text{BF}_3$  is a gas (m.p.  $-127^\circ\text{C}$ , b.p.  $-101^\circ\text{C}$ ) whereas boric acid,  $\text{B}(\text{OH})_3$ , is a solid that decomposes at  $185^\circ\text{C}$ .

Although dimethyl ether and ethyl alcohol both have the empirical formula  $\text{C}_2\text{H}_6\text{O}$ , ethyl alcohol boils at  $78.5^\circ\text{C}$  whereas dimethyl ether boils at  $-25^\circ\text{C}$ . The major portion of this difference is due to the extensive hydrogen bonding in ethyl alcohol that results in the formation of molecular aggregates that are more difficult to separate than are the polar dimethyl ether molecules that form no hydrogen bonds.

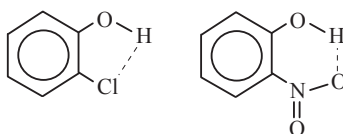
The effect of hydrogen bonding on the boiling points of glycerol and the ethyl ethers of glycerol are shown in the structures below. Replacing the OH groups with  $-\text{OC}_2\text{H}_5$  groups reduces the intermolecular hydrogen bonding, leading to lower boiling points.



Because of the intramolecular hydrogen bonding in the enol form of acetylacetone (2,4-pentadione), this compound tautomerizes to a great extent.



Likewise, the internal association of the hydroxyl hydrogen in the *ortho*-substituted phenols causes them to be less acidic than the *meta*- or *para*-isomers.

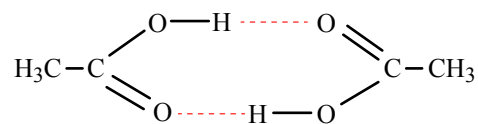


Hydrogen bonding affects many other types of physical and chemical properties. Liquids that are held together by sizable intermolecular forces often have heats of vaporization that are abnormally high. For many liquids that are held together by only van der Waals forces, the entropy of vaporization is essentially a constant, about  $88 \text{ J mol}^{-1} \text{ K}^{-1}$ .

$$\Delta S = S(g) - S(l) = \Delta H_{\text{vap}}/T \approx 88 \text{ J mol}^{-1} \text{ K}^{-1} \quad (4.16)$$

This relationship is known as *Trouton's rule*. If the liquid has significant structure as a result of intermolecular forces, it will have a lower entropy than a liquid that has a more random structure. Accordingly, the *change* in entropy when the liquid is converted into a completely random vapor will be larger. The larger  $\Delta S$  results from the low value for  $S(l)$  rather than a greater degree of randomness in the vapor. As a result, the entropy of vaporization is a useful parameter for interpreting the structure of liquids and vapors. Table 4.9 shows data for the vaporization of several liquids.

These data show clearly that liquids such as  $\text{NH}_3$ ,  $\text{CH}_3\text{OH}$ ,  $\text{HF}$ , and  $\text{H}_2\text{O}$  have higher entropies of vaporization because in the liquid state there is significant structure brought about by strong hydrogen bonding. The nonpolar liquids such as  $\text{CCl}_4$  and benzene are said to be “normal” or “unassociated” liquids. Acetic acid is a very different case although there is certainly strong hydrogen bonding in the liquid state. The very low entropy of vaporization is due to the fact that a random vapor is not produced. In fact, the vapor consists of dimers that can be represented as



Thus, the vapor has a lower entropy than it would have if the molecules were completely unassociated. Even though acetic acid has an entropy of vaporization that is much lower than that for other hydrogen-bonded liquids such as  $\text{CH}_3\text{OH}$  or water, the reason is again hydrogen bonding, but in the vapor phase as well as in the liquid phase. In effect, only one half a mole of independent species is present in the vapor when a mole of the liquid vaporizes.

3. *Spectral changes.* One of the most significant changes observed upon formation of hydrogen bonds by O–H or N–H groups is that produced in the infrared spectrum. As a result of the interaction, the O–H or N–H bond is slightly

**TABLE 4.9** Thermal Data for Vaporization of Liquids

Liquid	b.p., °C	$\Delta H_{\text{vap}}$ , (kJ mol <sup>-1</sup> )	$\Delta S$ (J mol <sup>-1</sup> K <sup>-1</sup> )
<i>n</i> -C <sub>4</sub> H <sub>10</sub>	-1.5	22.6	82
Cyclohexane	80.7	30.1	85
CCl <sub>4</sub>	76.1	30.0	86
SiH <sub>4</sub>	-112	10.0	83
SnH <sub>4</sub>	-51.8	12.1	86
GeH <sub>4</sub>	-90	14.1	77
Benzene	80.1	30.7	87
NH <sub>3</sub>	-33.4	23.3	97
N <sub>2</sub> H <sub>4</sub>	113.6	41.8	108
PH <sub>3</sub>	-87.8	14.6	79
NF <sub>3</sub>	-128.8	11.6	80
PF <sub>3</sub>	-101.5	16.5	96
SF <sub>4</sub>	-40.4	26.4	114
H <sub>2</sub> O	100.0	40.7	109
H <sub>2</sub> S	-59.6	18.7	87
CH <sub>3</sub> COOH	118.2	24.4	62
HF	19.5	30.3	104
CH <sub>3</sub> OH	64.7	35.3	104



**TABLE 4.10** Infrared Absorptions Associated with Hydrogen Bonding

IR region, $\text{cm}^{-1}$	Vibration	Assignment
3500–2500	X–H $\cdots$ :B / $\leftrightarrow$	$\nu_s$ X–H stretch
1700–1000	$\curvearrowright$ X–H $\cdots$ :B / $\curvearrowleft$	$\nu_b$ X–H in plane bend
400–300 (torsion)	$\uparrow$ X–H $\cdots$ :B / $\downarrow$	$\nu_t$ X–H out of plane bend
250–100	X–H $\cdots$ :B / $\leftrightarrow$	$\nu_{\text{H-B}}$ H–B stretching

weakened and lengthened, and the position of the band corresponding to the stretching vibration is shifted to lower wave numbers. This occurs regardless of whether the molecules undergo self-association or association with another electron donor.

In addition to the shift in the stretching band of the X–H bond, there are changes in the bending modes. As a result of “tying” the hydrogen atom to a pair of electrons on another atom, the stretching vibration occurs at a *lower* wave number than for the uncomplexed X–H bond. However, the two bending vibrations are hindered so that they are found at *higher* wave numbers for the hydrogen-bonded molecule. Table 4.10 shows the spectral effects that are the result of hydrogen bonding.

There have been numerous studies that have attempted to relate the extent of spectral shift of the stretching band  $\nu_s$  to the strength of the hydrogen bond when different donors are used. In general, a reasonably good correlation exists as long as the electron pair donor molecules have similar structures. For example, if the bases are all aliphatic amines, a relationship can be written as

$$-\Delta H = c\nu_s + b \quad (4.17)$$

where  $b$  and  $c$  are constants. If aromatic amines are used, they generally give a reasonably good relationship, but a different one than is found for aliphatic amines.

Hydrogen bonding is very important in determining the properties of both inorganic and organic materials. We will consider hydrogen bond formation as an acid–base interaction in Chapter 6. The treatment given here has barely scratched the surface of this important topic, and the interested reader should consult some of the books available on this subject.

## 4.2.5 Solubility Parameters

A property that is useful for interpreting intermolecular association in liquids is that known as the *solubility parameter*,  $\delta$ . When a liquid vaporizes, energy in the form of heat must be supplied to separate the molecules (to overcome the cohesion energy that holds the molecules together in the liquid state) and to perform the work done in expansion of the gas against atmospheric pressure. If the expansion work for 1 mol of gas is represented as  $RT$  (where  $R$  is  $8.314 \text{ J mol}^{-1} \text{ K}^{-1}$ ) and the heat of vaporization is given by  $\Delta H_{\text{vap}}$ , it follows that the cohesion energy of a mole of the liquid,  $E_c$  can be expressed by the equation

$$E_c = \Delta H_{\text{vap}} - RT \quad (4.18)$$

A quantity that is derived from the cohesion energy is known as the *specific cohesion* or the *cohesion density*,  $E_c/V$ , where  $V$  is the molar volume. The solubility parameter,  $\delta$ , is the square root of the specific cohesion, and it is expressed by the equation

$$\delta = (E_c/V)^{1/2} \quad (4.19)$$

Solubility parameters are extremely useful in predicting solubility and miscibility of liquids. For example, two liquids that have  $\delta$  values that differ significantly are usually not miscible. However, solubility parameters are also useful for interpreting molecular interactions. Values for solubility parameters typically range from about  $15 \text{ (J cm}^{-3}\text{)}^{1/2}$  for unassociated liquids such as  $\text{CCl}_4$  to about  $49 \text{ (J cm}^{-3}\text{)}^{1/2}$  for strongly associated liquids such as water. Instead of the units of  $\text{(J cm}^{-3}\text{)}^{1/2}$  or  $\text{J}^{1/2} \text{ cm}^{-3/2}$ , the unit  $\text{cal}^{1/2} \text{ cm}^{-3/2}$  is found in much of the literature and this unit is called a *hildebrand* in honor of Joel H. Hildebrand for his extensive work on liquids and solubility. The solubility parameters in the two sets of units differ by a factor of 2.045, the square root of the conversion factor from calories to joules, 4.184.

The cohesion energy of a liquid is inversely related to its vapor pressure. It is possible to derive a relationship between these properties that can be written as

$$E_c = \frac{d(\log P)}{d(1/T)} - RT \quad (4.20)$$

where  $P$  is the vapor pressure. It follows that, if one has an equation relating vapor pressure and temperature,  $E_c$  can be calculated from the  $P = f(T)$  equation. Although a large number of such equations exist, the *Antoine equation*,

$$\log P(\text{torr}) = A - \left( \frac{B}{C + t} \right) \quad (4.21)$$

(where  $A$ ,  $B$ , and  $C$  are constants and  $t$  is the temperature in  $^\circ\text{C}$ ), is one of the most convenient equations to use. Combining Eqs (4.20) and (4.21) we obtain

$$E_c = RT \left( \frac{2.303BT}{(C + t)^2} - 1 \right) \quad (4.22)$$

Fitting the vapor pressure data to the Antoine equation allows  $A$ ,  $B$ , and  $C$  to be determined. If the density of the liquid is available, the molar volume,  $V$ , is easily obtained. Then, the solubility parameter can be calculated using Eq. (4.19). Table 4.11 shows solubility parameters for a variety of liquids.

**TABLE 4.11** Solubility Parameters for Liquids

Compound	$\delta$ , $\text{J}^{1/2} \text{ cm}^{-3/2}$	Compound	$\delta$ , $\text{J}^{1/2} \text{ cm}^{-3/2}$
$\text{CF}_4$	17.0	$(\text{C}_2\text{H}_5)_4\text{Ge}$	17.6
$\text{CCl}_4$	17.6	$(n\text{-C}_3\text{H}_7)_4\text{Ge}$	18.0
$n\text{-C}_5\text{H}_{12}$	14.5	$(n\text{-C}_4\text{H}_9)_4\text{Ge}$	20.3
$n\text{-C}_8\text{H}_{18}$	15.3	$(n\text{-C}_5\text{H}_{11})_4\text{Ge}$	21.5
$\text{CH}_3\text{C}_6\text{H}_5$	18.2	$(\text{CH}_3)_3\text{Al}^a$	20.8
$\text{C}_6\text{H}_6$	18.8	$(\text{C}_2\text{H}_5)_3\text{Al}^a$	23.7
$\text{CS}_2$	20.5	$(n\text{-C}_3\text{H}_7)_3\text{Al}^a$	17.0
$\text{Br}_2$	23.5	$(i\text{-C}_4\text{H}_9)_3\text{Al}^a$	15.7
$\text{I}_2$	28.8	$\text{H}_2\text{O}$	47.9
$(\text{C}_2\text{H}_5)_3\text{B}$	15.4	$\text{CH}_3\text{COOH}$	21.3
$(\text{C}_2\text{H}_5)_2\text{Zn}$	18.2	$\text{CH}_3\text{OH}$	29.7
$\text{XeF}_2$	33.3	$\text{C}_2\text{H}_5\text{OH}$	26.6
$\text{XeF}_4$	30.9	$(\text{C}_2\text{H}_5)_2\text{O}$	15.8
$(\text{CH}_3)_4\text{Ge}$	13.6	$\text{CH}_3\text{COCH}_3$	20.0

<sup>a</sup>These compounds are dimers,  $(\text{R}_3\text{Al})_2$ .

Because molecular interaction can occur as a result of dipole–dipole forces, London dispersion forces, and hydrogen bonding, the overall cohesion energy,  $E_c$ , can be considered to be made up of contributions from each type of interaction (represented as  $E_D$ ,  $E_L$ , and  $E_H$ , respectively).

$$E_c = E_D + E_L + E_H \quad (4.23)$$

Dividing this equation by the molar volume,  $V$ , gives

$$\frac{E_c}{V} = \frac{E_D}{V} + \frac{E_L}{V} + \frac{E_H}{V} \quad (4.24)$$

Because the solubility parameter is given by  $\delta = (E_c/V)^{1/2}$ , it follows that

$$\delta^2 = \delta_D^2 + \delta_L^2 + \delta_H^2 \quad (4.25)$$

From this equation it can be seen that the solubility parameter is made up of contributions from each type of molecular force.

A complete description of the use of solubility parameters is beyond the scope of this book (for more details, see *Encyclopedia of Chemical Technology*, 2nd ed., Suppl. Vol., Interscience, New York, 1971, pp. 889–910). However, it is sufficient to point out here that liquids having  $\delta$  values that differ significantly tend to be immiscible or only partly soluble, but the *type* of intermolecular interaction is important in this regard (see Chapter 6). Also, note the high  $\delta$  values for liquids that hydrogen bond.

To illustrate the use and interpretation of solubility parameters, let us examine three cases. First, the alkyl germanes, which are nonpolar and are not associated in the liquid phase, show a regular, slight increase in  $\delta$  as the molecular masses, and hence the London forces, increase (see Table 4.11). This series represents one where all of the molecules interact by the same type of force with no tendency to dimerize. Second, the  $\delta$  values for ethanol and acetone, both of which have empirical formulas  $C_2H_6O$ , are 26.6 and 20.0  $J^{1/2} \text{ cm}^{-3/2}$ , respectively. The high value for the ethanol reflects intermolecular hydrogen bonding whereas acetone molecules interact only by weaker dipole–dipole and London forces.

A third and more subtle case is that of acetic acid. Although dimers result from hydrogen bonding in the liquid, it appears that the value of 21.3  $J^{1/2} \text{ cm}^{-3/2}$  is remarkably low compared to that for acetone (20.0  $J^{1/2} \text{ cm}^{-3/2}$ ). However, after the dimers of acetic acid are formed, there are rather weaker forces between these “units” so that there is no real tendency to form higher aggregates. For acetic acid, even vaporization leaves the dimers intact. Therefore, it is not surprising that a liquid composed of these dimers has a lower solubility parameter, even though strong hydrogen bonding is involved because of the very weak forces between the aggregates. An additional use of solubility parameters will be described in Chapter 10 to explain the association of aluminum alkyls.

Although the use of solubility parameters is extensive for organic compounds, this approach to understanding and interpreting intermolecular forces has received almost no attention in inorganic chemistry.

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In addition to a great amount of information on bonding, it also contains an enormous amount of information on solids and crystals.

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Smart, L., & Moore, E. (2005). *Solid State Chemistry* (3rd ed.). Boca Raton, FL: CRC Press. A readable introductory book that describes a great deal of behavior of inorganic solids.

Smith, D. W. (1990). *Inorganic Substances*. Cambridge: Cambridge University Press. A good discussion of energetics in solids is included in Chapter 5.

Weller, M. T. (1994). *Inorganic Materials Chemistry*. New York: Oxford University Press. A basic treatment of structure and properties of solids.

## PROBLEMS

- The H–H bond energy is  $431 \text{ kJ mol}^{-1}$ , the heat of formation of LiH is  $-90.4 \text{ kJ mol}^{-1}$ , and the lattice energy of LiH is  $916 \text{ kJ mol}^{-1}$ . Use these data and those shown in Table 4.1 to construct a thermochemical cycle and determine the electron affinity of the hydrogen atom.
- Consider a solid  $\text{MX}(s)$  composed of  $+1$  and  $-1$  ions. The process of dissolution can be considered as converting the crystal into gaseous ions followed by the interaction of the ions with the solvent as they become solvated. Write a thermochemical cycle to calculate the heat of solution in terms of the lattice energy and the solvation enthalpies of the ions.
- For a certain ionic compound, MX, the lattice energy is  $1220 \text{ kJ mol}^{-1}$  and the heat of solution in water is  $-90 \text{ kJ mol}^{-1}$ . If the heat of hydration of the cation is 1.50 times that of the anion, what are the heats of hydration of the ions? Use the thermochemical cycle developed in question 2.
- The radii of  $\text{Ag}^+$  and  $\text{Cl}^-$  are 126 and 181 pm, respectively. Use the Kapustinskii equation to calculate the lattice energy of AgCl.
  - The actual lattice energy for AgCl is  $912 \text{ kJ mol}^{-1}$ . Explain the difference between this value and the value calculated in part (a).
- Suppose you could make an aqueous solution containing  $\text{Na}^+$ ,  $\text{Ag}^+$ ,  $\text{F}^-$ , and  $\text{I}^-$  without a precipitate forming. If the solution were evaporated to dryness, what crystals would form? Use calculations from the Kapustinskii equation to support your conclusions.
- What crystal structure would you predict for a crystal containing  $\text{Fe}^{2+}$  and  $\text{F}^-$ ? The radii are 75 and 136 pm, respectively.
- Consider the crystal structure of  $\text{LiAlO}_2$  in which each Al is surrounded by six oxide ions each of which is bound to two aluminum atoms. If each Li is surrounded by four oxide ions, how many Li ions surround each oxide ion?
- Consider the sulfate ion,  $\text{SO}_4^{2-}$ , to be made up of  $\text{S}^{6+}$  and  $\text{O}^{2-}$  ions even though it is actually more covalent.
  - What would be the electrostatic bond character of the S–O bond?
  - If we consider solid  $\text{CaSO}_4$ , how many oxygen atoms must surround each  $\text{Ca}^{2+}$ ?
  - Having determined the number of bonds, what would the overall crystal structure of  $\text{CaSO}_4$  be? The radius of  $\text{Ca}^{2+}$  is 94 pm whereas that of  $\text{SO}_4^{2-}$  is 230 pm.
- Although  $\text{CaF}_2$  has the fluorite structure,  $\text{MgF}_2$  has the rutile structure. Explain this difference between these compounds.
- The O–H stretching band for  $\text{CH}_3\text{OH}$  is found at  $3649 \text{ cm}^{-1}$  when the alcohol is dissolved in heptane but it is found at  $3626 \text{ cm}^{-1}$  when the solvent is  $\text{CS}_2$ . Explain these observations.
- When  $\text{CH}_3\text{OH}$  is present in low concentration in  $\text{CCl}_4$ , a single sharp band is seen in the infrared spectrum at  $3642 \text{ cm}^{-1}$ . When a small amount of 4-cyanopyridine ( $\text{NCC}_5\text{H}_5\text{N}$ ) is added to the solution, two new bands appear at  $3430$  and  $3580 \text{ cm}^{-1}$ . Explain what these bands indicate.
- Explain why the acidity of  $o\text{-NO}_2\text{C}_6\text{H}_4\text{OH}$  is considerably different from that of  $p\text{-NO}_2\text{C}_6\text{H}_4\text{OH}$ .
- The boiling points of methanol and cyclohexane are  $64.7$  and  $80.7$  °C and the heats of vaporization are  $35,270$  and  $30,083 \text{ J mol}^{-1}$ , respectively. Determine the entropies of vaporization and explain the difference.
- The molecular masses of  $\text{CH}_3\text{NH}_2$  and  $\text{CH}_3\text{F}$  are approximately equal but the boiling points are  $-6.5$  and  $-78.4$  °C, respectively. Explain this difference.
  - The heats of vaporization of the two compounds given in (a) are  $16,680$  and  $27,070 \text{ J mol}^{-1}$ . Match these values to the appropriate compounds and test the validity of Trouton's rule for these compounds.

15. The process  $\text{O}(g) + 2e^- \rightarrow \text{O}^{2-}(g)$  absorbs  $653 \text{ kJ mol}^{-1}$ . In view of this, explain why so many ionic oxides exist.
16. Would you expect a higher melting point for CaO or  $\text{CaCl}_2$ ? Explain your answer.
17. What is the electrostatic bond character of the bonds in the  $\text{ReO}_3$  structure?
18. Explain the following observations related to the solubility of alkali halides.
  - (a) Lithium fluoride is almost insoluble in water, but the solubility increases for the series LiF, LiCl, LiBr, and LiI (the last of these being very soluble).
  - (b) Rubidium fluoride and iodide are very soluble in water, but the chloride and bromide are significantly less soluble.

# Reactions and Energy Relationships

In a study of inorganic chemistry, the number of chemical reactions encountered is enormous. If it were necessary to memorize all of these equations, the task would be formidable. However, most of the reactions discussed in this book can be grouped into a relatively small number of reaction *types*. Therefore, understanding these reaction types simplifies learning descriptive inorganic chemistry so a brief survey of several common types of reactions is presented in this chapter. Moreover, the energy changes accompanying reactions are important in understanding descriptive chemistry, and it is frequently possible to make predictions about possible reactions and stability of materials by the application of elementary thermodynamic principles. Because of this, we will begin this chapter with a brief review of some useful principles and applications of elementary thermodynamics.

## 5.1 THERMODYNAMIC CONSIDERATIONS

There are numerous physical transformations that can be considered as a system changing from one energy state to another. Chemical reactions also involve reactants and products that have different energies. As a result, it is important to understand the relationship between equilibrium and energy.

### 5.1.1 The Boltzmann Distribution Law

The solid, liquid, and vapor states of a substance represent the material in different energy states. Therefore, conversion of a substance from one state to another is accompanied by absorption or liberation of energy. The formation of a solution from a solvent and solute also involves an energy change that is known as the enthalpy of solution, and the enthalpy change for such processes may be positive or negative depending on the nature of the components of the solution. Figure 5.1 shows energy changes for some of these processes represented in energy level diagrams.

Situations such as those shown in Figure 5.1 exist throughout chemistry. In addition to the cases shown, the change of a solid from one crystal structure to another and the population of vibrational and rotational states of molecules can be represented as situations involving states of different energy. Because many types of behavior can be considered in terms of the population of states of unequal energy, it is important to have a way to relate the population of the states to the energy separating them. In other words, a principle that governs how the substances being studied will be “distributed” in the energy states is needed.

An explanation of how the population of states varies with the energy separating them is provided by a principle known as the Boltzmann distribution law which is named after Ludwig Boltzmann (1844–1906). A simplified form of the relationship can be written as

$$\frac{n_2}{n_1} = e^{-\Delta E/RT} \quad (5.1)$$

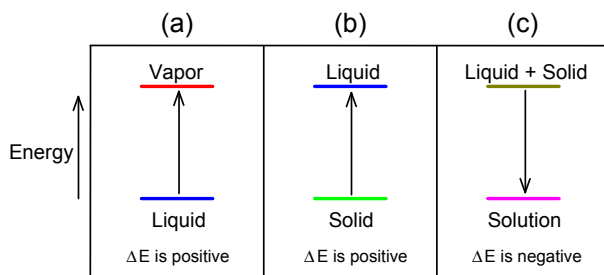
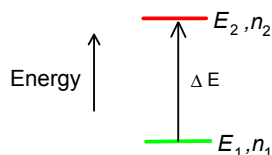


FIGURE 5.1 Some systems that can be considered as states of different energy.



**FIGURE 5.2** States of different energy populated according to a Boltzmann distribution where  $n_1$  and  $n_2$  are the populations of states having energies  $E_1$  and  $E_2$ , respectively.

where  $n_2$  and  $n_1$  are the populations of states having energies  $E_2$  and  $E_1$ ,  $\Delta E$  is the difference in energy between the two states,  $R$  is the molar gas constant ( $1.9872 \text{ cal mol}^{-1} \text{ K}^{-1}$  or  $8.3144 \text{ J mol}^{-1} \text{ K}^{-1}$ ), and  $T$  is the temperature (K). The application of this relationship can be illustrated by considering the populations of the energy states shown in Figure 5.2.

Let us illustrate the use of the Boltzmann distribution law by calculating the relative population of a sample that can exist in two energy states. For example, such a case might be the population of two vibrational energy levels for some type of molecule. The Boltzmann distribution law given in Eq. (5.1) allows us to calculate the relative population  $n_2/n_1$  if the temperature and the difference in energy between the states ( $\Delta E$ ) are known. Let us assume that the temperature involved is about room temperature, 300 K, and that the difference in energy between the states is  $2500 \text{ J mol}^{-1}$  ( $\Delta E = 2.5 \text{ kJ mol}^{-1}$ ). The value of  $RT$  at 300 K is ( $8.3144 \text{ J mol}^{-1} \text{ K}^{-1} \times 300 \text{ K}$ ), and that is approximately  $2500 \text{ J mol}^{-1}$ . Therefore, substituting these values in Eq. (5.1) gives

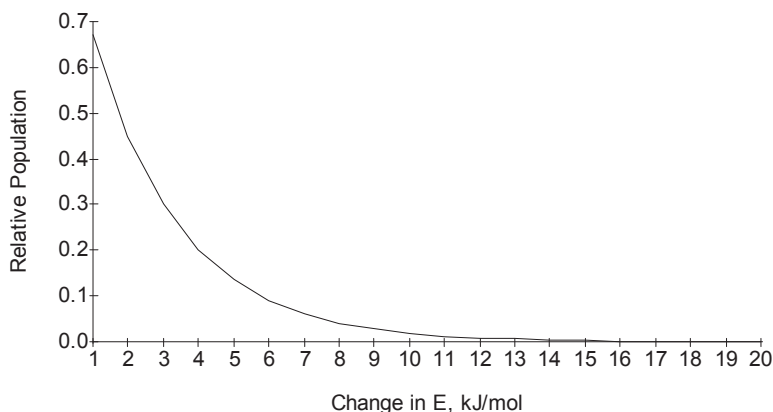
$$\frac{n_2}{n_1} = \exp(-2500 \text{ J mol}^{-1} / 2500 \text{ J mol}^{-1}) = e^{-1} = \frac{1}{e} = \frac{1}{2.718} = 0.368$$

It can be seen that the relative population  $n_2/n_1$  is 0.368 so only 36.8% as many molecules will populate the state having energy  $E_2$  as will populate the state having energy  $E_1$ . If the two states have the same energy,  $\Delta E = 0$ , and  $n_2/n_1 = e^0 = 1$ . This indicates that the two states would be equally populated. Such a result is reasonable because the two states have the same energy (they are degenerate states) and, therefore, occupancy in either is equally probable. If the energy difference is about  $5000 \text{ J mol}^{-1}$ , then

$$\frac{n_2}{n_1} = \exp(-5000 \text{ J mol}^{-1} / 2500 \text{ J mol}^{-1}) = e^{-2} = 0.135$$

This result shows that in this case the population of the higher energy state is only 13.5% that of the lower one. It is apparent that  $n_2/n_1 = 1$  if the states have equal energy and that the relative population ( $n_2/n_1$ ) decreases exponentially as  $\Delta E$  increases if the temperature remains constant. Figure 5.3 shows a graph of the relative population of the states at 300 K as a function of the energy difference between them. Even for states separated by only a few kilojoules, the relative population of the state having higher energy is very small.

It can be seen that the relative population of the higher energy state decreases as the energy difference increases. However, because  $T$  occurs in the denominator of the exponential function, one way to increase the population of the upper state is to increase the temperature. Going from the lower state to the upper one is an endothermic process, and increasing the temperature favors the endothermic direction of the process.



**FIGURE 5.3** Population of a state of higher energy relative to the ground state at 300 K.

An interesting manipulation of Eq. (5.1) is to solve for  $n_2$  and take the natural logarithm of both sides of the resulting equation to obtain

$$\ln n_2 = -\frac{\Delta E}{RT} + \ln n_1 \quad (5.2)$$

If the population of the state of lower energy is much larger than that of the upper one (which is often the case), the value of  $n_1$  is essentially a constant. Such a situation exists, for example, when a small amount of liquid evaporates in a closed container leaving a considerable amount of the liquid phase. Therefore, we can write Eq. (5.2) as

$$\ln n_2 = -\frac{\Delta E}{RT} + \text{Constant} \quad (5.3)$$

or in a linear form,

$$\ln n_2 = -\frac{\Delta E}{R} \left( \frac{1}{T} \right) + \text{Constant} \quad (5.4)$$

$$y = m(x) + b$$

The population of the vapor state determines the vapor pressure of a liquid. As long as the vapor and the liquid are in equilibrium, it does not matter how much liquid is present. The difference in “energy” between the liquid and vapor states is the heat of vaporization,  $\Delta H_{\text{vap}}$ , so we can write

$$\ln p = -\frac{\Delta H_{\text{vap}}}{R} \left( \frac{1}{T} \right) + \text{Constant} \quad (5.5)$$

This equation shows that the logarithm of the vapor pressure,  $p$ , plotted versus  $1/T$  should give a straight line having a slope of  $-\Delta H_{\text{vap}}/R$ . This is, in fact, one way to determine the heat of vaporization of a liquid. In this case, the slope of the line is negative so that  $\Delta H_{\text{vap}}$  is positive, indicating that heat is *absorbed* in vaporizing a liquid. Of course, when a vapor condenses to a liquid the heat of vaporization is *evolved* ( $\Delta H_{\text{cond}}$ ) and has a negative value ( $-\Delta H_{\text{vap}}$ ).

In applying these principles to solubility, it should be remembered that the excess of a solid solute can be considered to be a constant. Therefore, interpreting Figure 5.1(c) using this approach suggests that when the solubility of a substance has been determined at several temperatures, a plot of the  $\ln(\text{solubility})$  of the substance versus  $1/T$  should yield a straight line that can be used to determine the heat of solution of the compound. This is one way in which the heat of solution can be obtained, but it is also possible to determine the heat of solution by direct calorimetric measurements.

The discussion above serves to illustrate that many properties governed by the Boltzmann distribution law are related to temperature in such a way that the natural logarithm of the property plotted versus  $1/T$  will yield a straight line having a slope that is related to the heat change associated with the property. However, if the heat change is itself a function of temperature, the relationship is only approximate. In most cases, the heat change associated with a chemical or physical process can be treated as a constant as long as the temperature range is not very great. Therefore, using the relationship between  $\ln(\text{some property})$  and  $1/T$  provides a way to determine the heat change accompanying the process.

For a chemical reaction, there are three possible ways in which the energy can change, and these are illustrated in Figure 5.4. The term “reaction energy” is not quite appropriate. In the case of a chemical reaction, we need a more precise thermodynamic definition. The thermodynamic quantities that are important for chemical reactions are  $\Delta H$ ,  $\Delta S$ , and  $\Delta G$ . If the reader has not been introduced to basic thermodynamics, a general chemistry text or a thermodynamics book should be consulted.

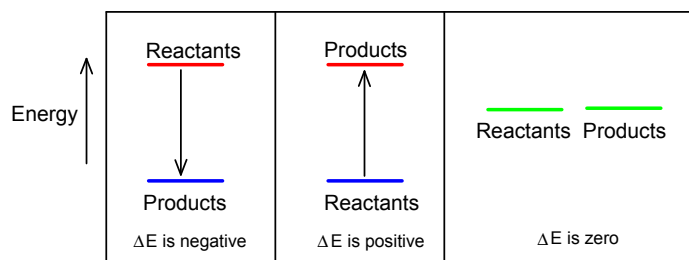


FIGURE 5.4 Three possible energy changes accompanying a chemical reaction.



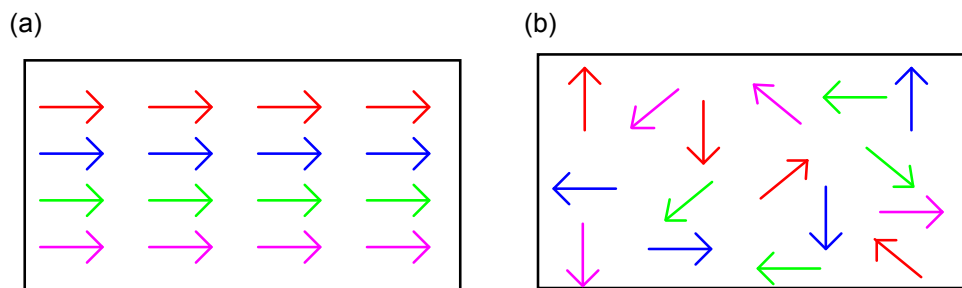


FIGURE 5.5 Systems that illustrate low entropy (a) and high entropy (b).

The enthalpy change,  $\Delta H$ , is the change in heat associated with a process. However, the concept of *entropy* is based on the randomness or disorder of a system, and it can be illustrated by reference to Figure 5.5. A system that is highly ordered (structured and having low entropy) is shown in Figure 5.5(a) whereas a system having a high degree of disorder (unstructured and having high entropy) is shown in Figure 5.5(b). By analogy to these simple illustrations, we would expect a crystalline solid to have low entropy (because it has a regular structure), a somewhat ordered liquid to have higher entropy, and a gas with randomly distributed molecules to have still higher entropy because entropy is a measure of randomness or disorder of the system. Therefore, the change in entropy,  $\Delta S$ , for a physical or chemical transformation is associated with the change in the disorder of the system. The change in *Gibbs free energy*,  $\Delta G$ , is related to the changes in enthalpy and entropy by the equation

$$\Delta G = \Delta H - T\Delta S \quad (5.6)$$

Some elementary but important uses of this equation will now be illustrated.

### 5.1.2 Reactions and $\Delta G^\circ$

A fundamental thermodynamic principle can be stated in simple terms as *water flows downhill*. Systems spontaneously progress to a state of lower energy. For a chemical reaction, the energy in question is  $\Delta G$ , and  $\Delta G$  must be *negative* for the reaction to be spontaneous. Note that nothing is being said about the *rate* of the process, and that topic will be addressed later. Figure 5.4 should now be modified to relate reactions to  $\Delta G$  rather than energy. Figure 5.6 shows  $\Delta G$  for three possible types of reactions.

It has already been stated that the criterion for a spontaneous chemical reaction is that  $\Delta G$  must be negative. Let us consider a reaction that can be represented as follows.



The reaction when A and B are initially mixed takes place toward the right but eventually the rate of the reaction to the left becomes equal to the rate of the reaction to the right. At that point, equilibrium is established. The equilibrium constant for the reaction shown in Eq. (5.7) can be written as

$$K = \frac{[C][D]}{[A][B]} \quad (5.8)$$

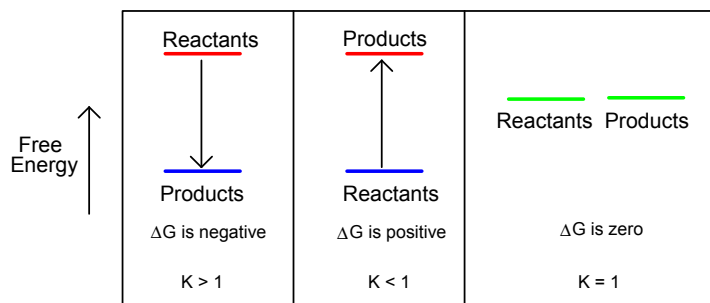


FIGURE 5.6 Three possible free energy changes that may accompany a reaction.

where [ ] means molar concentration of the species contained within the brackets. If the equation has balancing coefficients other than 1, as in the reaction



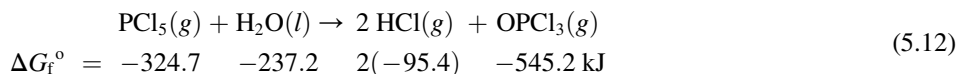
the coefficients appear as exponents in the equilibrium constant expression. Although we will not derive the relationship here,  $\Delta G$  is related to the equilibrium constant by the equation

$$\Delta G = -RT \ln K \quad (5.10)$$

It can be seen that if  $\Delta G = 0$  then  $K = 1$ , and the reaction proceeds neither to the right nor the left preferentially. Because  $-\Delta G/RT = \ln K$ , we can see that the more negative  $\Delta G$  is, the larger the value of  $\ln K$  (and hence  $K$ ). Therefore, reactions for which  $\Delta G$  has a large negative value proceed farther to the right (to form products). It follows, then, that if we could determine  $\Delta G$  we would know theoretically whether the reaction in question could take place or not. This would not, however, tell us anything about the rate of the reaction. Even reactions which are energetically favorable may take place slowly because there may be no low energy pathway. For a chemical reaction, the overall  $\Delta G^\circ$  is simply the sum of the free energies of formation of the products minus the sum of the free energies of formation of the reactants,

$$\Delta G^\circ = \sum \Delta G_f^\circ(\text{products}) - \sum \Delta G_f^\circ(\text{reactants}) \quad (5.11)$$

Because large compilations of  $\Delta G_f^\circ$  values are available, it is often possible to determine easily whether a given reaction will “work.” For example, consider the reaction shown below where the free energy of formation of each substance is written below its formula.

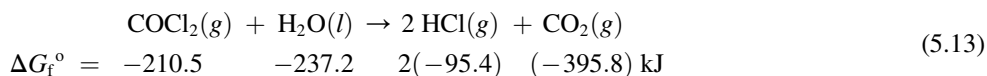


Thus, for the reaction shown in the equation above, it can be seen that

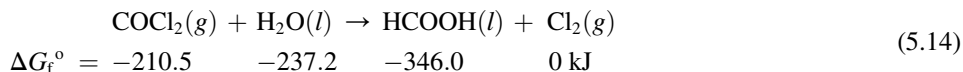
$$\Delta G^\circ = [-545.2 + 2(-95.4)] - [(-324.7) + (-237.2)] = -174.1 \text{ kJ}$$

and the reaction is spontaneously possible because  $\Delta G^\circ$  is negative. It is important to realize that a reaction that does not take place spontaneously at one set of conditions may become favorable at a different set of conditions (temperature, pressure, etc.).

As illustrated by the following examples, thermodynamics provides a powerful tool for predicting the feasibility of a reaction. Consider the two possible reactions that are shown below for carbonyl chloride (phosgene) reacting with water.

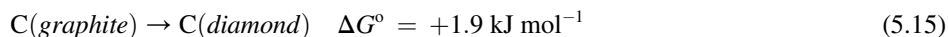


In this case,  $\Delta G^\circ = -138.9 \text{ kJ}$  so it is possible for this reaction to take place with the release of energy. Using the same reactants, we might also suppose that another reaction between  $\text{COCl}_2$  and  $\text{H}_2\text{O}$  could be represented as



This reaction leads to  $\Delta G^\circ = +101.7 \text{ kJ}$  so is not likely that the reaction will take place in this way. At first glance the two equations look might equally plausible, but a very simple application of thermodynamic principles allows us to predict that the reaction of  $\text{COCl}_2$  with water produces  $\text{HCl}$  and  $\text{CO}_2$  rather than  $\text{HCOOH}$  and  $\text{Cl}_2$ . It should be noted that from a chemical point of view, it would be very unlikely that a strong oxidizing agent such as  $\text{Cl}_2$  would be produced by a reaction involving water (which not normally an oxidizing agent).

It must be emphasized again that even reactions that are “unlikely” are not necessarily impossible, especially under different conditions. Changing the conditions (i.e., temperature and pressure) may cause the reaction to become feasible because  $\Delta G$  may be negative under a different set of conditions. This is especially true for reactions where  $\Delta G$  is only slightly positive. For example, the transformation of graphite to diamond can be shown as



This suggests that diamond could not be prepared from graphite. However, the transformation can be carried out (and is on a large scale) under special conditions (see Chapter 11).

### 5.1.3 Relationship between $\Delta G$ and $T$

To determine how  $\Delta G^\circ$  is related to  $T$  we will assume that  $\Delta H^\circ$  and  $\Delta S^\circ$  for a reaction are constant even though the temperature varies. This is generally a valid assumption as long as the temperature does not change greatly. At a certain temperature,  $T_1$ , the free energy change is

$$\Delta G_1^\circ = \Delta H^\circ - T_1 \Delta S^\circ \quad (5.16)$$

But as we have seen,  $\Delta G^\circ$  is also related to the equilibrium constant as shown in Eq. (5.10),

$$\Delta G_1^\circ = -RT_1 \ln K_1$$

where  $K_1$  is the equilibrium constant at  $T_1$ . Therefore,

$$\Delta H^\circ - T_1 \Delta S^\circ = -RT_1 \ln K_1 \quad (5.17)$$

so that

$$T_1 \Delta S^\circ = RT_1 \ln K_1 + \Delta H^\circ \quad (5.18)$$

If we divide both sides of Eq. (5.18) by  $T_1$ , we obtain

$$\Delta S^\circ = R \ln K_1 + \frac{\Delta H^\circ}{T_1} \quad (5.19)$$

We obtain an analogous equation for  $\Delta S^\circ$  at a different temperature  $T_2$ ,

$$\Delta S^\circ = R \ln K_2 + \frac{\Delta H^\circ}{T_2} \quad (5.20)$$

Equating the two expressions for  $\Delta S^\circ$ , we obtain

$$R \ln K_2 + \frac{\Delta H^\circ}{T_2} = R \ln K_1 + \frac{\Delta H^\circ}{T_1} \quad (5.21)$$

which can be simplified to yield

$$\ln \frac{K_2}{K_1} = \frac{\Delta H^\circ}{R} \left( \frac{1}{T_1} - \frac{1}{T_2} \right) = \frac{\Delta H(T_2 - T_1)}{RT_1 T_2} \quad (5.22)$$

Of course, for a series of temperatures ( $T_i$ ) Eq. (5.17) can be written as

$$\Delta H^\circ - T_i \Delta S^\circ = -RT_i \ln K_i \quad (5.23)$$

or

$$\ln K_i = -\frac{\Delta H^\circ}{R} \left( \frac{1}{T_i} \right) + \frac{\Delta S^\circ}{R} \quad (5.24)$$

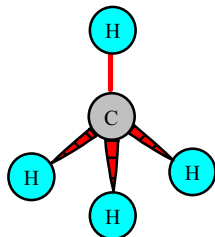
But because  $\Delta S^\circ$  can normally be considered to be constant for a particular reaction, Eq. (5.24) can be written as

$$\ln K_i = -\frac{\Delta H^\circ}{R} \left( \frac{1}{T_i} \right) + \text{Constant} \quad (5.25)$$

This relationship shows that determining the equilibrium constant at a series of temperatures  $T_i$  and plotting  $\ln K_i$  versus  $1/T_i$  should yield a straight line having a slope  $-\Delta H/R$  thus allowing the reaction  $\Delta H$  to be determined. Once again it is observed that there is a linear relationship between the natural logarithm of some property and  $1/T$ .

### 5.1.4 Bond Enthalpies

If we consider the molecule CH<sub>4</sub>, the structure shows that it has four equivalent C–H bonds.



Therefore, separating the molecule into gaseous atomic species as shown in the following equation



requires an enthalpy which is four times the *average* enthalpy of a C–H bond. Methane has a heat of formation of  $-78.85 \text{ kJ mol}^{-1}$  whereas that of gaseous carbon is  $+718.4 \text{ kJ mol}^{-1}$ , and that of gaseous hydrogen atoms is  $+217.9 \text{ kJ mol}^{-1}$ . Therefore, the reaction shown in Eq. (5.26) has an enthalpy change at 298 K of

$$\begin{aligned} \Delta H^\circ &= \sum \Delta H_f^\circ(\text{products}) - \sum \Delta H_f^\circ(\text{reactants}) \\ &= [4(217.9) + 718.4] - [(-78.85) - 4(0.008314)(298)] \\ &= 1679 \text{ kJ} \end{aligned} \quad (5.27)$$

As a result of five moles of gaseous product having been produced from 1 mol of gaseous reactant, there is a net increase of 4 mol of gaseous materials, and the last term in Eq. (5.27) represents the work done as 4 mol of gas expand or push back the atmosphere. That amount of work is expressed as the change in  $PV$  which can be written as  $4 \times d(PV)$ . If the pressure is constant,  $4 \times d(PV) = 4 \times P(dV) = 4RT$ . In most cases, this term can be ignored unless the change in number of moles of gaseous reactants and products is rather large.

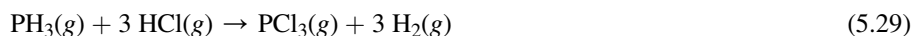
The actual value of 1679 kJ is the total dissociation enthalpy of 4 C–H bonds which gives an average enthalpy of  $419.9 \text{ kJ mol}^{-1}$  per bond. Removing the H atoms one at a time requires a slightly different energy for each because the first H atom comes from CH<sub>4</sub>, the second from CH<sub>3</sub>, etc. If other molecules containing C–H bonds are included, it is possible to obtain an average C–H bond enthalpy which is about  $414 \text{ kJ mol}^{-1}$ .

If enthalpies of atomization of molecules containing bonds between atoms of numerous types are considered, it is possible to obtain values for the bond enthalpies for a large variety of bonds. As we have seen, bond enthalpies are very useful in predicting reactions because it is quite possible that the necessary free energies and/or enthalpies of formation may not be readily available for all of the reactants and products. Bond enthalpies are usually given as *positive* values as if the bonds are being *broken*. If the bonds are being *formed*, the *negative* values are used. Therefore, for a chemical reaction, some bonds are broken (heat *absorbed*) and others are formed (heat *released*). It is easy to show that

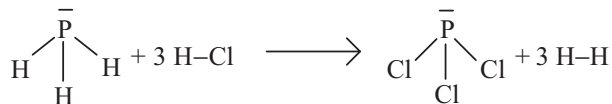
$$\Delta H(\text{reaction}) = \sum \Delta H(\text{bonds formed}) + \sum \Delta H(\text{bonds broken}) \quad (5.28)$$

It must be remembered that mathematically bonds formed give a *negative* value and bonds broken give a *positive* value. In order to make use of this approach, bond enthalpies are needed, and Table 5.1 shows average bond enthalpies for numerous kinds of bonds.

If we consider the reaction written as



bond enthalpies can be used to predict whether or not the reaction is possible. The structures of all of the reactants and products make it possible to determine the number of bonds of each type that are involved in the reaction.



**TABLE 5.1** Average Bond Enthalpies

Bond	$\text{kJ mol}^{-1}$	Bond	$\text{kJ mol}^{-1}$
H–H	435	N=O	607
H–F	565	N–H	389
H–Cl	430	F–F	155
H–Br	362	Cl–Cl	243
H–I	297	Br–Br	193
H–S	339	I–I	151
H–P	318	O–O	138
H–Si	318	O=O <sup>a</sup>	494
H–Li	243	S–S	213
H–Na	197	S=S	423
C–C	347	P–P	222
C=C	619	P=P	485
C≡C	812	Si–Si	188
C–O	335	Li–Li	108
C=O	707	Na–Na	74
C–N	293	K–K	49
C=N	616	F–Cl	252
C≡N	879	Cl–Br	218
C–Cl	326	Cl–I	208
C–Br	285	Cl–S	272
C–H	414	Cl–P	322
N–N	159	Cl–Si	264
N=N	418	Cl–O	205
N≡N	941	O–S	272
N–O	201	O=S	507
N–Br	163	C=S	477
H–O	464	C–F	485

<sup>a</sup>Bond energy in the O<sub>2</sub> molecule in which the bond order is 2.

For this reaction, the summary of bond changes is as follows

Bonds Broken: (Positive)	Bonds Formed: (Negative)
3 P–H = $3 \times 318$ kJ	3 P–Cl = $3 \times (-322)$ kJ
3 H–Cl = $3 \times 430$ kJ	3 H–H = $3 \times (-435)$ kJ
Total = +2244 kJ	Total = –2271 kJ

Because the enthalpy released by forming the bonds in the products is greater than that absorbed by breaking bonds in the reactants,  $\Delta H$  is negative for this reaction ( $\Delta H = -27$  kJ for the reaction as written). However, the criterion of spontaneity is that  $\Delta G$  should be negative, and both  $\Delta H$  and  $\Delta S$  are necessary to determine  $\Delta G$ . Is it possible to make any predictions about this reaction by making use of bond enthalpies? Note that as shown in Eq. (5.29), there

are 4 mol of gaseous reactants and 4 mol of gaseous products. Therefore, the total number of moles of gaseous substances does not change. Because entropies of gases are much larger than those of solids or liquids, it is the gaseous materials that concern us most in dealing with  $\Delta S$  for a reaction. In the present case,  $\Delta S$  will be very small because of there being equal numbers of moles of gaseous reactants and products. Therefore, the  $-T\Delta S$  term in the equation giving  $\Delta G$  will be small compared to  $\Delta H$  which means that  $\Delta G$  will be approximately equal to  $\Delta H$ . This is indicated by the equation

$$\Delta G = \Delta H - T\Delta S \approx \Delta H \quad (5.30)$$

It can be shown that the actual  $\Delta S$  value for the reaction shown in Eq. (5.29) is only  $0.066 \text{ kJ K}^{-1}$  so we are correct in our assumption that  $\Delta G$  is approximately the same as  $\Delta H$ , and our prediction of spontaneity based on  $\Delta H$  is valid in this case.

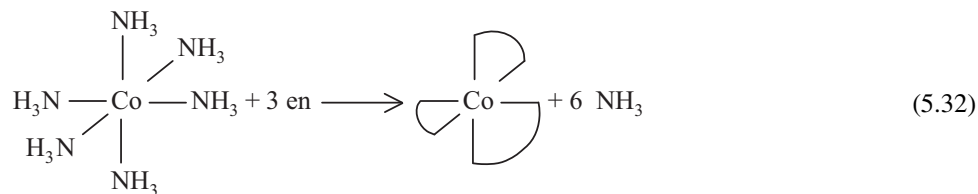
If a reaction has an enthalpy change of about zero, the reaction may still be spontaneous because a *negative* value for  $\Delta G$  can result from a *positive* value for  $\Delta S$  as shown by the equation

$$\Delta G = \Delta H - T\Delta S = 0 - T\Delta S$$

It can be seen that when  $\Delta H \approx 0$ ,  $\Delta G$  will be negative if  $\Delta S$  is positive. Consider for example, the reaction



where en represents  $\text{H}_2\text{NCH}_2\text{CH}_2\text{NH}_2$ , ethylenediamine, a molecule that has two nitrogen atoms that bond to a  $\text{Co}^{3+}$  ion simultaneously. The bonds between nitrogen atoms and the cobalt ion have about the same strength regardless of whether the nitrogen atom is contained in  $\text{NH}_3$  or a  $\text{H}_2\text{NCH}_2\text{CH}_2\text{NH}_2$  molecule. Therefore, the *enthalpy* change for the reaction is very small because six  $\text{Co}^{3+}-\text{N}$  bonds having approximately the same strength are involved in either case. However, when we consider the complex formed,  $\text{Co}(\text{en})_3^{3+}$ , it is seen that each en molecule bonds to the  $\text{Co}^{3+}$  in two places. The arrangement of the bonds is illustrated in the following equation from which it can be seen that the  $\text{Co}^{3+}$  ion forms six bonds pointed toward the corners of an octahedron.



Because each en molecule has two atoms that can function as electron pair donors to the  $\text{Co}^{3+}$ , only *three* ethylenediamine molecules are required to replace *six*  $\text{NH}_3$  molecules to complete the coordination sphere of  $\text{Co}^{3+}$ . Therefore, for the reaction shown in Eq. (5.32), there is a net increase in the number of unbound (free or randomly arranged) molecules because six  $\text{NH}_3$  molecules are replaced by only three ethylenediamine molecules. As a result, the reaction leads to an increase in disorder and  $\Delta S$  has a rather large *positive* value that results in  $\Delta G$  being negative even though  $\Delta H$  is approximately zero. Therefore, the equilibrium constant for the reaction is large, and the reaction takes place spontaneously.

When writing equations, two sets of possible products may look reasonable until the reaction is considered from the point of view of the thermodynamic changes involved. In many cases, the use of thermodynamic data allows the correct products to be identified. The predictive power of thermodynamics in deciding how reactions take place should not be overlooked.

## 5.2 COMBINATION REACTIONS

Numerous reactions of the elements can be classified as combination reactions because two substances combine to form a product. Reactions such as those given below are, of course, redox reactions because changes in oxidation states also occur.





Generally, a nonmetal is the oxidizing agent although the reducing agent may be either a metal or a nonmetal. Many elements can exist in more than one positive oxidation state, and this leads to the possibility of multiple products even for reactions of the same two elements. For example,  $\text{CO}_2$  and  $\text{CO}$  can result for the combustion of carbon.



As a practical rule, it is usually assumed that if an excess of the oxidizing agent is present, the element being oxidized will go to its highest available oxidation state. Thus, when an excess of oxygen is present, C reacts to form  $\text{CO}_2$ . When there is a deficiency of oxygen (excess of C),  $\text{CO}$  is formed. There is an important exception to this rule and it involves sulfur. Regardless of how much oxygen is available, the reaction leads to  $\text{SO}_2$  rather than  $\text{SO}_3$  unless a catalyst is used.

There are other exceptions with regard to the products formed as in the case of reactions of Group IA metals with oxygen. Lithium gives the expected product,  $\text{Li}_2\text{O}$ .



However, the remaining elements of the group give other oxides (see Chapter 8) as illustrated by the following equations.



This behavior is also shown by the Group IIA metals to some extent. It is also useful to remember that nitrogen is somewhat unreactive for a nonmetal having a high electronegativity (3.0). This results from the stability of the nitrogen molecule,  $\text{N}\equiv\text{N}$ , for which the bond energy is  $941 \text{ kJ mol}^{-1}$ . It is also useful to remember that although the reaction of sulfur with fluorine can be shown as



the analogous reaction with chlorine



does not occur. Doubtless this is due to at least two factors:

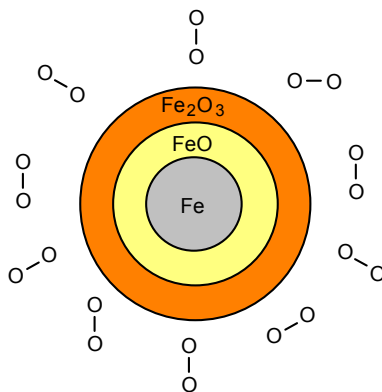
1.  $\text{Cl}_2$  is not nearly as strong an oxidizing agent as  $\text{F}_2$ .
2. It is not possible to get six atoms as large as Cl around S without some repulsion.

As a result of these factors,  $\text{SCl}_6$  is not stable but  $\text{SF}_6$  is.

The reaction shown in Eq. (5.36) deserves additional discussion because this process is not nearly as simple as the equation indicates. Consider an iron object surrounded by oxygen gas and reacting as illustrated in Figure 5.7. The oxygen reacts on the surface of the object where  $\text{Fe}_2\text{O}_3$  is formed (the product showing the highest ratio of oxygen to iron). Below this layer, there is still unreacted iron or products containing a smaller ratio of oxygen to iron. As the reaction proceeds, iron atoms must diffuse through the  $\text{Fe}_2\text{O}_3$  product layer, and below that outer layer there is an excess of Fe present for the  $\text{O}_2$  to react with. Thus,  $\text{FeO}$  is formed initially, but below that layer the object is still unreacted Fe. At the interface of layers of  $\text{Fe}_2\text{O}_3$  and  $\text{FeO}$ , an oxide of intermediate composition,  $\text{Fe}_2\text{O}_3 \cdot \text{FeO}$  or  $\text{Fe}_3\text{O}_4$ , is present. After a considerable time, most of the iron will have been converted to  $\text{Fe}_2\text{O}_3$ . However, some unreacted iron and some  $\text{FeO}$  are still likely to be present. At some point, the product is considered to be  $\text{Fe}_2\text{O}_3$  although some  $\text{FeO}$  (and possibly some Fe) may still remain. Although we write the reaction of Fe or other metals with oxygen (or some other gas) as producing the expected product with the metal in the highest oxidation state, the reaction is never that simple, and the product is unlikely to be a single compound having exact stoichiometry.

In many cases, two compounds react so there is no reason not to consider reactions such as





**FIGURE 5.7** An illustration of the reaction of gaseous oxygen with an iron object.

as combination reactions. Certainly the reaction shown in Eq. (5.44) is the combination of two compounds to form a third. However, it is also an acid–base reaction in the Lewis sense (see Chapter 6), and these reactions are probably better described in these terms rather than as combination reactions.

### 5.3 DECOMPOSITION REACTIONS

Many compounds are not highly stable. Therefore, they undergo a variety of decomposition reactions when they are heated, exposed to light ( $h\nu$ ), or receive some other form of energy. The following are examples of such reactions.



Also, some compounds decompose into two or more products that may or may not have been the starting materials for the preparation of the original compound. For example,  $\text{NH}_4\text{Cl}$  can be prepared by the reaction of  $\text{NH}_3$  and  $\text{HCl}$ , and when heated,  $\text{NH}_4\text{Cl}$  decomposes to produce these compounds as products.



Similarly, heating  $\text{CaCO}_3$  produces lime,  $\text{CaO}$ , and  $\text{CO}_2$ .



In both of these reactions the products can recombine to reform the starting materials. However, some decomposition reactions lead to products that will not react to reform the original material. Examples of two such compounds are  $\text{NH}_4\text{NO}_2$  and  $\text{NH}_4\text{NO}_3$  that decompose as shown in the following equations.

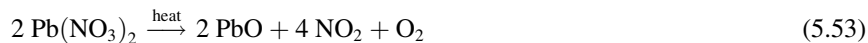


These reactions, especially the latter, are potentially dangerous because  $\text{NH}_4\text{NO}_3$  can explode if it is heated above approximately  $200^\circ\text{C}$ . Such a situation exists when an element (N in this case) is present in two different oxidation states, one positive and the other negative. In this case,  $\text{NO}_3^-$  is a strong oxidizing agent and  $\text{NH}_4^+$  is a reducing agent. In neither of the reactions shown in Eqs (5.50) and (5.51) do the products react to reform the starting materials.

Decomposition of compounds that may look as if they should be chemically similar can produce greatly different products. Consider, for example, the heating of two nitrates and two azides.







Although the decomposition reactions discussed so far involve heating, the absorption of energy in other forms can also lead to decomposition of compounds. Light has an energy given by  $h\nu$  where  $h$  is Planck's constant and  $\nu$  is the frequency of the light. When light strikes silver bromide, it decomposes, and  $\text{Fe}(\text{CO})_5$  decomposes when subjected to ultrasound.



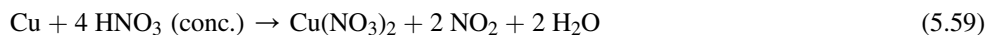
These cases show that decomposition reactions of many types have been studied. We will see many more examples of decomposition reactions in the remainder of this book.

## 5.4 REDOX REACTIONS

It is appropriate to consider the combination reaction of carbon and oxygen,



as an oxidation–reduction or redox reaction. Carbon increases in oxidation number from 0 to +4 (oxidation), and the oxygen atoms change from 0 to –2 (reduction). However, there are numerous other redox reactions that cannot be classified as combinations or decompositions. One way of classifying some of these redox reactions is as *chemical* redox reactions in which a chemical substance is the oxidizing agent. Examples of this type include the following cases in which  $\text{HNO}_3$  is the oxidizing agent.



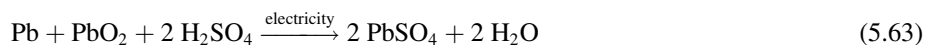
These reactions illustrate an important principle of redox chemistry which is that the actual products may depend on the relative concentration of the oxidizing and/or reducing agents. This was previously illustrated for certain of the combination reactions as well. The reaction



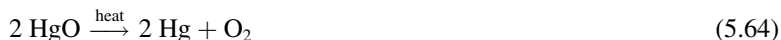
is another example of the chemical redox type that can be used for the laboratory preparation of chlorine.

Although we will study the chemistry of many oxidizing and reducing agents in subsequent chapters, you should recall that strong oxidizing agents include  $\text{NO}_3^-$ ,  $\text{Cr}_2\text{O}_7^{2-}$ ,  $\text{MnO}_4^-$ ,  $\text{H}_2\text{O}_2$ ,  $\text{OCl}^-$ ,  $\text{F}_2$ ,  $\text{Cl}_2$ , and  $\text{ClO}_4^-$ . These species undergo reduction when they act as oxidizing agents. The actual fate of these species in a redox reaction depends on the reducing agent and the experimental conditions. For example, in acidic solutions  $\text{MnO}_4^-$  is usually reduced to  $\text{Mn}^{2+}$ , but in basic solutions  $\text{MnO}_2$  is the usual reduction product. Common reducing agents include  $\text{H}_2$ , metals, carbon, sulfur, phosphorus, and numerous other species.

Other redox reactions are called *electrochemical* redox reactions because they either consume or produce electricity. Examples of this type are the following.



Finally, some reactions can be considered as *thermal* redox reactions because they take place only at high temperatures. Equations (5.64–5.66) represent reactions of this type.



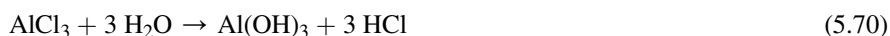
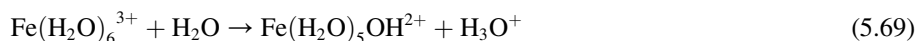


## 5.5 HYDROLYSIS REACTIONS

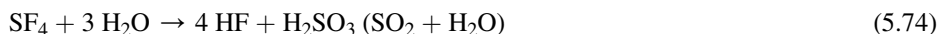
There are numerous important reactions in which a molecule of water reacts in such a way that part of the molecule (H) appears in one product and the remainder (OH) appears in another. In other words, a molecule of water is split or *lysed* which leads to the name *hydrolysis* that describes these reactions. Typical processes include reactions such as those that follow indicating that a base is produced.



However, there are also hydrolysis reactions that produce acidic solutions as shown in the following equations.



Characteristically, the covalent bonds between nonmetals and halogens are very susceptible to hydrolysis reactions. A few examples of this behavior are illustrated in the following equations, and numerous others will be seen in later chapters dealing with the chemistry of the nonmetallic elements.



In general, reactions such as these produce the hydrogen halide and an acid containing the nonmetal in the same oxidation state as the original halogen compound. One notable exception to this behavior is that of  $\text{SF}_6$  because it does not react with water (see Chapter 16).

Analogous *lysis* reactions such as those above also take place in other solvents. For example, in liquid  $\text{NH}_3$  the reaction



is a *lysis* reaction. The less restrictive term which applies to solvents other than water is *solvolysis*. In the reaction above, the HCl initially liberated reacts with  $\text{NH}_3$  to produce  $\text{NH}_4\text{Cl}$ .

## 5.6 REPLACEMENT REACTIONS

In many reactions, one atom or group of atoms replaces another. For example, zinc replaces copper from a solution of copper sulfate.



The reverse reaction does not occur, and the prediction of reactions such as this is based on the redox potentials (*electromotive force*,  $\mathcal{E}^\circ$ ). We have already seen that the criterion for spontaneity of a reaction depends on the sign of  $\Delta G^\circ$ . Therefore, there must be a relationship between  $\mathcal{E}^\circ$  and  $\Delta G^\circ$  and it is

$$\Delta G^\circ = -n\mathcal{F}\mathcal{E}^\circ = -RT \ln K \quad (5.77)$$

where  $n$  is the number of electrons transferred,  $\mathcal{F}$  is Faraday's constant (96,000 coulombs) and  $\mathcal{E}^\circ$  is the voltage. From Eq. (5.77) we can see that  $\Delta G^\circ$  will be negative if  $\mathcal{E}^\circ$  is positive. Consult a general chemistry text if you are unclear

on how to determine  $\mathcal{E}^\circ$  for a reaction. For the reaction shown in Eq. (5.76), it is found that  $\mathcal{E}^\circ = +1.10$  V. From Eq. (5.77) we find that

$$-RT \ln K = -n\mathcal{F}\mathcal{E}^\circ \quad (5.78)$$

and solving for  $\ln K$  we obtain

$$\ln K = \frac{n\mathcal{F}\mathcal{E}^\circ}{RT} \quad (5.79)$$

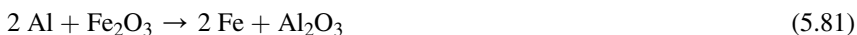
Substituting in the appropriate values for the quantities when  $T = 298$  K gives

$$\mathcal{F}/RT = 38.75$$

and

$$\ln K = 2 \times 1.10 \times 38.75 = 85.6$$

Therefore,  $K = e^{85.6} = 1.50 \times 10^{37}$  indicating that the equilibrium in the reaction for zinc replacing copper lies far to the right. The reactions



also represent replacement reactions where one metal replaces another. Although the  $\mathcal{E}^\circ$  must be calculated for each reaction, the general trend is that a metal of higher electronegativity liberates one of lower electronegativity. Of course, reactions such as



in which an active metal replaces hydrogen readily occur.

For nonmetals, replacement reactions also take place.



This reaction occurs because  $\text{Cl}_2$  is a stronger oxidizing agent than  $\text{Br}_2$ . In general, a stronger oxidizing agent liberates a weaker one. Therefore,  $\text{F}_2$  will oxidize  $\text{Cl}^-$  to produce  $\text{Cl}_2$ ;  $\text{Cl}_2$  will oxidize  $\text{Br}^-$  to produce  $\text{Br}_2$ ; etc.

Other types of replacement reactions occur in which a group of atoms or a molecule is displaced. For example,  $\text{SO}_3$  displaces  $\text{CO}_2$  in the reaction

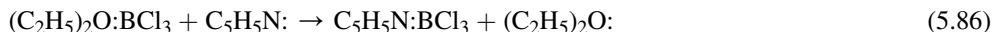


This reaction occurs because the stronger Lewis acid,  $\text{SO}_3$ , displaces the weaker one,  $\text{CO}_2$ . If we consider  $\text{CO}_3^{2-}$  to be an oxide ion bound to  $\text{CO}_2$ , the affinity of  $\text{SO}_3$  for the  $\text{O}^{2-}$  is greater so  $\text{SO}_3$  displaces  $\text{CO}_2$  and forms  $\text{SO}_4^{2-}$ . A strong acid generally displaces a weaker one from its compounds. Similarly, a strong base will displace weaker bases from their compounds. In the following reactions, consider  $\text{NH}_4^+$  to be  $\text{NH}_3$  attached to  $\text{H}^+$ . The hydroxide ion has a greater affinity for  $\text{H}^+$  than does  $\text{NH}_3$  so there is a transfer of  $\text{H}^+$ .



This reaction occurs because  $\text{OH}^-$  is a stronger base than  $\text{NH}_3$ , and this reaction is often used as a qualitative test for ammonium salts. These acid–base reactions will be treated more fully in Chapter 6.

Reactions such as



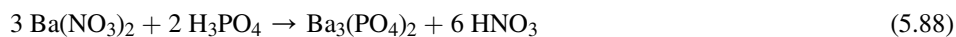
are also replacement reactions as a result of the stronger Lewis base,  $\text{C}_5\text{H}_5\text{N}$  (pyridine), replacing the weaker one,  $(\text{C}_2\text{H}_5)_2\text{O}$  (ether). This type of acid–base chemistry is discussed in greater detail in Chapter 6.

## 5.7 METATHESIS

Reactions of the type



and



are called metathesis or exchange reactions. However, when solutions of NaCl and KNO<sub>3</sub> are mixed, no visible change occurs because the potential products KCl and NaNO<sub>3</sub> are also soluble in water.



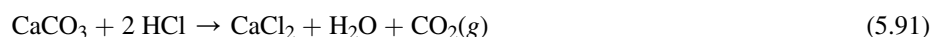
Therefore, in order for a reaction of this type to take place, a product must effectively be removed from the reaction zone by one of the following processes.

1. A precipitate is formed.



In this case, BaSO<sub>4</sub> is insoluble.

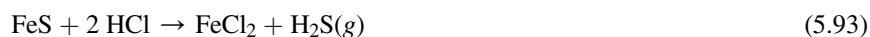
2. A gas is formed.



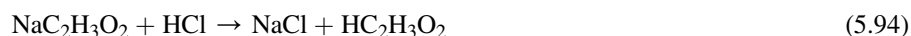
Carbonic acid, H<sub>2</sub>CO<sub>3</sub>, is unstable so it decomposes and liberates CO<sub>2</sub>. A gas is also evolved when sulfites react with acids,



because the expected product from the exchange, H<sub>2</sub>SO<sub>3</sub>, is unstable. When treated with an acid, sulfides also evolve a gas, H<sub>2</sub>S.



3. A slightly ionized product is formed.



In this case, acetic acid is essentially unionized because it is a weak acid (see Chapter 6). The neutralization reaction



is of this type because water is an unionized substance.

4. A complex is formed.

Many metal ions, especially those of transition metals, have an affinity for groups that donate electron pairs. This results in the formation of a complex in which the metal ion binds to anions, H<sub>2</sub>O, NH<sub>3</sub>, and other species because these species have unshared pairs of electrons. In the following reaction, Pt<sup>2+</sup> forms bonds to four Cl<sup>-</sup> ions to produce [PtCl<sub>4</sub>]<sup>2-</sup>.



Generally, Pt<sup>2+</sup> forms stable complexes in which it is bonded to four ions or molecules rather than some other number. This type of chemistry will be discussed in more detail in Chapters 20 and 21.

In order to know if a precipitate will form as a product of a reaction in aqueous solution, it is necessary to know a few simple *solubility rules*. These are given in all standard general chemistry texts, but they will be given here to aid in writing equations.

1. All nitrates, acetates, and chlorates are soluble.
2. All chlorides and bromides are soluble except those of Ag<sup>+</sup>, Pb<sup>2+</sup>, and Hg<sub>2</sub><sup>2+</sup>.
3. All iodides are soluble except Ag<sup>+</sup>, Hg<sub>2</sub><sup>2+</sup>, Cu<sup>+</sup>, Pb<sup>2+</sup>, and Hg<sup>2+</sup>.

4. All sulfates are soluble except  $\text{Ba}^{2+}$ ,  $\text{Sr}^{2+}$ ,  $\text{Pb}^{2+}$ ,  $\text{Hg}^{2+}$ , and  $\text{Ag}^+$ .  $\text{CaSO}_4$  is slightly soluble.
5. All sulfides are insoluble except Group IA, IIA, and  $\text{NH}_4^+$ .
6. All carbonates are insoluble except Group IA, IIA,  $\text{NH}_4^+$ .
7. All phosphates are insoluble except Group IA and  $\text{NH}_4^+$ .
8. All hydroxides are insoluble except Group IA. Barium, strontium, and calcium hydroxides are slightly soluble.

With regard to the formation of a gaseous product, it should also be remembered that sulfides, carbonates, sulfites, and nitrites evolve a gas when treated with an acid (see the discussion above). Ammonium salts evolve  $\text{NH}_3$  when treated with a strong base (see Eq. (5.85)).

Slightly ionized (or unionized) materials include water, weak acids (acetic, hydrofluoric, phosphoric, nitrous, phosphorous, organic acids, etc.), and weak bases (particularly “ $\text{NH}_4\text{OH}$ ” and organic bases such as amines). Complexes are formed by a wide range of species, particularly transition metal ions and molecules or ions that have unshared pairs of electrons (see Chapters 6 and 20).

## 5.8 NEUTRALIZATION REACTIONS

The reaction of an acid with a base to the destruction of the properties of both is a neutralization, and  $\text{H}_2\text{O}$  (an unionized product) results. Thus, the following equations represent neutralization reactions because the acidic and basic properties of the reactants are removed.



As a result of the ionization of acids and bases in water, the essential reaction in each case can be shown as

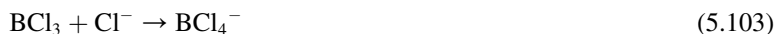


In the Brønsted description of acids and bases, *an acid is a proton donor and a base is a proton acceptor*. Therefore, the reaction



is an acid–base reaction because there is proton transfer (see Chapter 6) even though the reactants are not dissolved in water.

According to the Lewis theory of acids and bases (to be discussed in detail in Chapter 6), *an acid is an electron pair acceptor and a base is an electron pair donor*. Accordingly, the following reactions are acid–base reactions because they represent processes in which electron pair donation and acceptance occurs.



In the first two of these reactions,  $\text{BCl}_3$  behaves as a Lewis acid because the boron atom has only three pairs of electrons surrounding it in the  $\text{BCl}_3$  molecule so it functions as an electron pair acceptor. Typical of most ions of transition metals,  $\text{Cu}^{2+}$  readily accepts electron pairs from  $\text{NH}_3$  molecules.

Because of the many types of interactions that can be classified as acid–base interactions, an enormous number of reactions can be considered as acid–base processes. A large number of reactions of this type will be illustrated in the following chapters. Although the brief survey presented in this chapter does not show all of the types of reactions that will be encountered in the study of inorganic chemistry, the majority of the reactions shown in later chapters are of these types. As you study the remaining chapters, try to classify the reactions you see as belonging to the types discussed. In that way, numerous reactions will become additional examples of these types and not just isolated cases to be memorized. Moreover, a thorough understanding of the reaction types will enable you to predict the outcome of many reactions, and this will make learning descriptive chemistry more than just memorizing a great number of equations.

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## PROBLEMS

1. What is the relative population of two states that differ in energy by 25 kJ at a temperature of 300 K?
2. The solubility of  $\text{KBrO}_3$  in water (given as grams dissolved per 100 g of water) varies with temperature as follows:

Temp., °C	10	20	30	50	60	80
Solubility	4.8	6.9	9.5	17.5	22.7	34.0

Use these data to determine the heat of solution of  $\text{KBrO}_3$  in water ( $R = 8.3144 \text{ J/mol K}$ ).

3. The vapor pressure of  $\text{CCl}_4$  as a function of temperature is as follows:

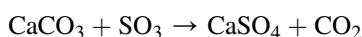
Temp., °C	30	35	40	45	50	55	60
V.P., torr	143	176	216	262	317	379	451

Use these data to determine the heat of vaporization of  $\text{CCl}_4$  ( $R = 8.3144 \text{ J/mol K}$ ).

4. Given the following  $\Delta G_f^\circ$  values, decide whether the preparation of sulfuryl chloride,  $\text{SO}_2\text{Cl}_2$ , from  $\text{PCl}_5$  and  $\text{SO}_3$  is feasible.

$\text{PCl}_5(\text{g}) - 305 \text{ kJ/mol}$	$\text{SO}_2\text{Cl}_2(\text{g}) - 314 \text{ kJ/mol}$
$\text{SO}_3(\text{g}) - 371 \text{ kJ/mol}$	$\text{OPCl}_3(\text{l}) - 521 \text{ kJ/mol}$

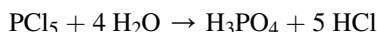
5. For the following reaction



determine whether the reaction is feasible, given the following  $\Delta G_f^\circ$  values.

$\text{CaCO}_3(\text{s}) - 1129 \text{ kJ/mol}$	$\text{CaSO}_4(\text{s}) - 1332 \text{ kJ/mol}$
$\text{SO}_3(\text{g}) - 371 \text{ kJ/mol}$	$\text{CO}_2(\text{g}) - 394 \text{ kJ/mol}$

6. For the following reaction

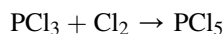


determine whether the reaction is feasible, given the following  $\Delta G_f^\circ$  values.

$\text{PCl}_5(\text{g}) - 305 \text{ kJ/mol}$	$\text{H}_3\text{PO}_4(\text{s}) - 1119 \text{ kJ/mol}$
$\text{H}_2\text{O}(\text{l}) - 237 \text{ kJ/mol}$	$\text{HCl}(\text{g}) - 95 \text{ kJ/mol}$

7. Using the table of bond energies, calculate  $\Delta H$  values for the following reactions.
- $2 \text{PH}_3 \rightarrow \text{P}_2\text{H}_2 + 2 \text{H}_2$
  - $3 \text{H}_2 + \text{N}_2 \rightarrow 2 \text{NH}_3$
  - $4 \text{ClF}_3 + \text{CH}_4 \rightarrow 4 \text{ClF} + \text{CF}_4 + 4 \text{HF}$
  - $2 \text{H}_2\text{S} + \text{CO}_2 \rightarrow 2 \text{H}_2\text{O} + \text{CS}_2$
8. Using bond enthalpies, determine which of the following reactions (not balanced) are likely to be spontaneous.
- $\text{F}_2(\text{g}) + \text{CH}_4(\text{g}) \rightarrow \text{HF}(\text{g}) + \text{CF}_4(\text{g})$
  - $\text{PH}_3(\text{g}) + \text{HCl}(\text{g}) \rightarrow \text{PCl}_3(\text{g}) + \text{H}_2(\text{g})$
  - $\text{H}_2\text{S}_2(\text{g}) + \text{CCl}_4(\text{g}) \rightarrow \text{S}_2\text{Cl}_2(\text{g}) + \text{CH}_4(\text{g})$
  - $\text{SiCl}_4 + \text{H}_2\text{O} \rightarrow \text{SiH}_4 + \text{HOCl}$
9. Using the table of bond enthalpies, determine which of the following reactions (not balanced) would likely take place as written.
- $\text{P}_2\text{H}_4(\text{g}) + \text{Cl}_2(\text{g}) \rightarrow \text{PCl}_3(\text{g}) + \text{HCl}(\text{g})$
  - $\text{HCN}(\text{g}) + \text{H}_2(\text{g}) \rightarrow \text{CH}_3\text{NH}_2(\text{g})$
  - $\text{HI}(\text{g}) + \text{Cl}_2(\text{g}) \rightarrow \text{ICl}_3(\text{g}) + \text{HCl}(\text{g})$
  - $\text{NH}_3(\text{g}) + \text{Br}_2(\text{g}) \rightarrow \text{NBr}_3(\text{g}) + 3 \text{HBr}(\text{g})$
10. Using the table of bond energies, determine the approximate enthalpy change for the following reactions.
- $2 \text{PCl}_3 \rightarrow \text{P}_2\text{Cl}_4 + \text{Cl}_2$
  - $2 \text{CH}_3\text{OH} + 3 \text{O}_2 \rightarrow 2 \text{CO}_2 + 4 \text{H}_2\text{O}$
  - $\text{H}_2\text{C}=\text{CH}_2 + \text{H}_2 \rightarrow \text{C}_2\text{H}_6$
  - $\text{N}_2 + 2 \text{H}_2 \rightarrow \text{N}_2\text{H}_4$
11. Complete and balance the following:
- $\text{S}_8 + \text{F}_2$  (excess)  $\rightarrow$
  - $\text{Cs} + \text{O}_2 \rightarrow$
  - $\text{P}_4 + \text{S}_8$  (excess)  $\rightarrow$
  - $\text{As} + \text{Cl}_2$  (excess)  $\rightarrow$
  - $\text{Fe} + \text{Cl}_2$  (excess)  $\rightarrow$
12. Complete and balance the following. Note any special conditions.
- $\text{Mg} + \text{N}_2 \rightarrow$
  - $\text{Fe} + \text{Cl}_2$  (excess)  $\rightarrow$
  - $\text{Ca} + \text{H}_2 \rightarrow$
  - $\text{Cu} + \text{S}_8$  (excess)  $\rightarrow$
  - $\text{B} + \text{S}_8 \rightarrow$
13. Complete and balance the following if a reaction occurs. Assume all reactants are in aqueous solutions. If a reaction occurs, tell why.
- $\text{K}_2\text{SO}_3 + \text{H}_2\text{SO}_4 \rightarrow$
  - $\text{CuCl}_2 + \text{KOH} \rightarrow$
  - $\text{CaCl}_2 + \text{NH}_4\text{NO}_3 \rightarrow$
  - $\text{Fe}(\text{OH})_3 + \text{HCl} \rightarrow$
  - $\text{NaNO}_3 + (\text{NH}_4)_3\text{PO}_4 \rightarrow$
  - $(\text{NH}_4)_2\text{S} + \text{HCl} \rightarrow$
14. Complete and balance the following if a reaction occurs. Assume all reactants are in aqueous solutions. If a reaction occurs, tell why.
- $\text{NH}_4\text{Cl} + \text{KOH} \rightarrow$
  - $\text{Na}_2\text{SO}_4 + \text{Ba}(\text{NO}_3)_2 \rightarrow$
  - $\text{K}_2\text{S} + \text{LiCl} \rightarrow$
  - $\text{PCl}_3 + \text{H}_2\text{O} \rightarrow$
  - $\text{FeSO}_4 + \text{K}_2\text{CO}_3 \rightarrow$

15. Complete and balance the following if a reaction occurs. Assume all reactants are in aqueous solutions. If a reaction occurs, tell why.
- $\text{K}_2\text{SO}_3 + \text{HCl} \rightarrow$
  - $\text{LiC}_2\text{H}_3\text{O}_2 + \text{HCl} \rightarrow$
  - $\text{Na}_2\text{S} + \text{ZnCl}_2 \rightarrow$
  - $\text{Na}_3\text{PO}_4 + \text{H}_2\text{SO}_4 \rightarrow$
  - $\text{CuCO}_3 + \text{HCl} \rightarrow$
16. Predict the direction in which each of the reactions should be more favorable.
- $\text{CaO} + \text{MgS} \rightleftharpoons \text{CaS} + \text{MgO}$
  - $\text{RbI} + \text{NaF} \rightleftharpoons \text{RbF} + \text{NaI}$
  - $(\text{C}_2\text{H}_5)_2\text{O}:\text{BCl}_3 + \text{N}(\text{CH}_3)_3 \rightleftharpoons (\text{CH}_3)_3\text{N}:\text{BCl}_3 + (\text{C}_2\text{H}_5)_2\text{O}$
  - $\text{Al}_2\text{O}_3 + 3 \text{Ni} \rightleftharpoons 2 \text{Al} + 3 \text{NiO}$
  - $\text{NaC}_2\text{H}_3\text{O}_2 + \text{HCl} \rightleftharpoons \text{NaCl} + \text{HC}_2\text{H}_3\text{O}_2$
17. Write equations for the following reactions carried out at high temperature.
- $\text{Ca} + \text{NH}_4\text{Cl}(l) \rightarrow$
  - $\text{CaCO}_3 + \text{P}_4\text{O}_{10} \rightarrow$
  - $\text{Na}_2\text{CO}_3 + \text{SiO}_2 \rightarrow$
  - $\text{AlF}_3 + \text{NaF} \rightarrow$
  - $\text{MgO} + \text{SO}_3 \rightarrow$
18. Write equations for the reactions of molten  $\text{NH}_4\text{Cl}$  (an acid) with each of the following.
- Mg
  - CaO
  - $\text{Fe}_2\text{O}_3$
  - CaS
19.  $\text{TiO}_2$  is a stronger acidic oxide than is  $\text{CO}_2$ . Write the equation for the reaction that would occur between  $\text{BaCO}_3$  and  $\text{TiO}_2$  at high temperature. Describe the structure of the solid product.
20. A reaction in which an atom is *oxidized* at the same time that it *adds* additional groups is often referred to as an oxidation–addition reaction (abbreviated *oxad*). A reaction of this type is



in which P is oxidized from +3 to +5 as it forms two additional bonds to chlorine atoms. Complete the following equations and draw the structure of the product.

- $\text{SCl}_2 + \text{Cl}_2 \rightarrow$
- $\text{OSCl}_2 + \text{Cl}_2 \rightarrow$
- $\text{P}_4\text{O}_6 + \text{O}_2 \rightarrow$
- $\text{SnF}_2 + \text{F}_2 \rightarrow$
- $\text{ClF}_3 + \text{F}_2 \rightarrow$



# Acids, Bases, and Nonaqueous Solvents

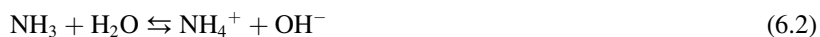
When one considers the incredible number of chemical reactions that are possible, it becomes apparent why a scheme that systemizes a large number of reactions is so important and useful. Indeed, classification of reaction types is important in all areas of chemistry, and a great deal of inorganic chemistry can be systematized or classified by the broad types of compounds known as acids and bases. Many properties and reactions of substances are understandable and predictions can often be made about their reactions in terms of acid–base theories. In this chapter, we will describe the most useful acid–base theories and show their applications to inorganic chemistry. However, water is not the only solvent that is important in inorganic chemistry, and a great deal of chemistry has been carried out in other solvents. In fact, the chemistry of nonaqueous solvents is currently a field of a substantial amount of research in inorganic chemistry so some of the fundamental nonaqueous solvent chemistry will be described in this chapter.

## 6.1 ACID–BASE CHEMISTRY

According to the *Arrhenius theory*, an acid was defined as any substance that produced the hydrogen ion,  $H^+$  in aqueous solution. We now write  $H^+$  as  $H_3O^+$  to show that it is solvated, but the actual species is probably  $H_9O_4^+$  ( $H^+$  surrounded by four  $H_2O$  molecules) in dilute solutions. A base was defined as any substance that produced  $OH^-$  in aqueous solution. Thus, HCl is an acid because the reaction



produces  $H_3O^+$ . Similarly,  $NH_3$  is a base because the reaction



produces  $OH^-$  ions. Neutralization can be represented by the equation



for combinations of acids and bases that dissociate in aqueous solutions. Of course, the anion from the acid and the cation from the base form a salt that is usually soluble in water. One difficulty with this theory is that it does not apply to gaseous reactions such as



where no solvent is present. Also, according to the Arrhenius theory the properties of an acid are limited to the properties of  $H_3O^+$  and those of a base are limited to the properties of  $OH^-$ . These ions are not the reacting species shown in Eq. (6.4), but it is no less an acid–base reaction. In this case, a salt,  $NH_4Cl$ , is formed, but there is no water formed.

The *Brønsted–Lowry theory* (usually called the *Brønsted theory*), advanced by these workers in 1923, is more comprehensive than the Arrhenius theory. According to this theory, an acid–base reaction is characterized as a reaction in which a proton is transferred from one species (the acid, the *proton donor*) to another (the base, the *proton acceptor*). There can be no acid without a base; neither exists in isolation because the proton must be transferred from the acid to some other species. According to this theory, HCl is an acid because when it is placed in water, it acts as a proton donor,



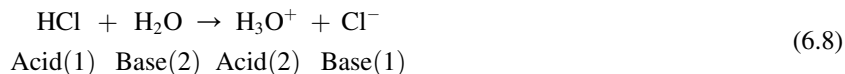
not because it produces  $H_3O^+$  ions. It is a natural consequence of this view that if  $H_2O$  accepts a proton, it is functioning as a base. After the  $H_2O$  molecule has accepted a proton, it then has the capacity to function as a proton donor to another species such as  $OH^-$ .



When one species functions as a proton donor, it always produces another species that can function as a potential proton donor that is weaker than the first. Accordingly,  $\text{H}_2\text{O}$  must be able to act as an acid, which it does in many instances as illustrated by the equation



A similar situation exists for bases. In Eq. (6.7), the very strong base  $\text{O}^{2-}$  has produced a weaker base,  $\text{OH}^-$ . The pairs of substances differing by the transfer of a proton (e.g.,  $\text{NH}_3$  and  $\text{NH}_4^+$  or  $\text{H}_2\text{O}$  and  $\text{OH}^-$ ) are said to be *conjugate pairs*. This is illustrated by Eq. (6.8).



In this equation,  $\text{Cl}^-$  is the conjugate base of  $\text{HCl}$ , and  $\text{H}_3\text{O}^+$  is the conjugate acid of  $\text{H}_2\text{O}$ .

Seeing that the Brønsted definitions of acids and bases are not restricted to a specific solvent, this theory can readily explain the reaction shown in Eq. (6.4) in which there is no solvent. In that case,  $\text{HCl}$  donates a proton to  $\text{NH}_3$  resulting in the formation of the ionic salt  $\text{NH}_4\text{Cl}$ . Therefore,  $\text{HCl}$  is a Brønsted acid. Because  $\text{NH}_3$  accepts a proton, it is acting as a base. Likewise, the Brønsted theory is applicable to many reactions in which there is a solvent other than water, which makes the Brønsted theory much more widely applicable than is the Arrhenius theory.

According to the Arrhenius theory, the strength of an acid depends on the fraction of the acid that ionizes to produce  $\text{H}_3\text{O}^+$  ions. When defined in terms of the Brønsted theory, the acid strength is reflected by the magnitude of the equilibrium constant for the *proton transfer* reaction



for which the equilibrium constant is

$$K = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}][\text{H}_2\text{O}]} \quad (6.10)$$

For several acids, e.g.,  $\text{HCl}$ ,  $\text{HBr}$ ,  $\text{HNO}_3$ ,  $\text{HClO}_4$ , or  $\text{H}_2\text{SO}_4$ , this reaction proceeds essentially to completion in dilute aqueous solutions. Therefore, all of these acids are completely dissociated, and all appear to have equal strength. This is known as the *leveling effect* in which water has caused all the acids to appear to have the same strength. This occurs because water is able to act as a base toward these strong acids. Therefore, in water solutions, acid strength is limited (or leveled) to that of  $\text{H}_3\text{O}^+$ . If a solvent is chosen that is not nearly so readily a proton acceptor, the acids are shown to have different strengths as proton donors. For example, in acetic acid, which is not normally a proton acceptor, the extent of the reaction



is somewhat different for the acids listed above. The equilibrium constant is largest when  $\text{HA}$  is  $\text{HClO}_4$ . Although acetic acid is normally a weak acid, relative to  $\text{HClO}_4$  it is a weak base. On the basis of similar equilibria for a series of strong acids, the strength as proton donors appears to decrease in the order  $\text{HClO}_4 > \text{HBr} > \text{H}_2\text{SO}_4 > \text{HCl} > \text{HNO}_3$ . Thus, even though these acids all *appear* to have equal strength in water, they actually have different strengths. It also follows from the Brønsted concept that the greater the strength of an acid, the weaker the strength of its conjugate as a base. Because  $\text{HClO}_4$  is such a strong acid, the  $\text{ClO}_4^-$  ion must be a very weak conjugate base.

The dissociation constant for an acid  $\text{HA}$  can be written as

$$K_a = K[\text{H}_2\text{O}] = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]} \quad (6.12)$$

To be more correct thermodynamically, *activities* (the activity of a species is its concentration multiplied by a number known as the activity coefficient) should be used instead of concentrations. However, in dilute solutions, the activity coefficients are usually taken to be 1 so concentrations are used instead. Table 6.1 shows the dissociation constants of several common acids, and a more complete list can be found in the *CRC Handbook of Chemistry and Physics*. Because the strength of an acid is reflected by the magnitude of the dissociation constant, the acids shown in the table are arranged in the order of decreasing acid strength. The leveling effect causes the first five to *appear* to be of equal strength in dilute aqueous solutions.

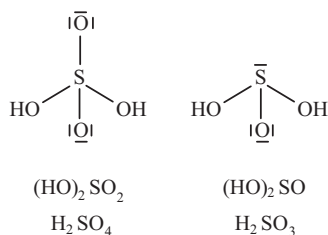
**TABLE 6.1** Dissociation Constants of Acids in Water

Acid	Dissociation Constant ( $K_a$ )	Conjugate Base
HClO <sub>4</sub>	Essentially complete	ClO <sub>4</sub> <sup>-</sup>
HBr	Essentially complete	Br <sup>-</sup>
HCl	Essentially complete	Cl <sup>-</sup>
HSCN	Essentially complete	SCN <sup>-</sup>
H <sub>2</sub> SO <sub>4</sub>	Essentially complete	HSO <sub>4</sub> <sup>-</sup>
H <sub>2</sub> C <sub>2</sub> O <sub>4</sub>	$5.9 \times 10^{-2}$	HC <sub>2</sub> O <sub>4</sub> <sup>-</sup>
HSO <sub>4</sub> <sup>-</sup>	$2.0 \times 10^{-2}$	SO <sub>4</sub> <sup>2-</sup>
HClO <sub>2</sub>	$1.0 \times 10^{-2}$	ClO <sub>2</sub> <sup>-</sup>
H <sub>3</sub> PO <sub>4</sub>	$7.5 \times 10^{-3}$	H <sub>2</sub> PO <sub>4</sub> <sup>-</sup>
H <sub>3</sub> AsO <sub>4</sub>	$4.8 \times 10^{-3}$	H <sub>2</sub> AsO <sub>4</sub> <sup>-</sup>
H <sub>2</sub> Te	$2.3 \times 10^{-3}$	HTe <sup>-</sup>
HF	$7.2 \times 10^{-4}$	F <sup>-</sup>
H <sub>2</sub> Se	$1.7 \times 10^{-4}$	HSe <sup>-</sup>
HN <sub>3</sub>	$1.9 \times 10^{-5}$	N <sub>3</sub> <sup>-</sup>
HTe <sup>-</sup>	$1.0 \times 10^{-5}$	Te <sup>2-</sup>
H <sub>2</sub> S	$9.1 \times 10^{-8}$	HS <sup>-</sup>
HSe <sup>-</sup>	$1.0 \times 10^{-10}$	Se <sup>2-</sup>
HS <sup>-</sup>	$1.2 \times 10^{-15}$	S <sup>2-</sup>
H <sub>2</sub> O	$1.1 \times 10^{-16}$	OH <sup>-</sup>

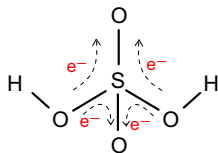
### 6.1.1 Factors Affecting Acid Strength

When one considers the series of acids HX (where X = F, Cl, Br, or I) and H<sub>2</sub>X (where X = O, S, Se, or Te), it is seen that there is a general increase in acid strength with increasing size of the element attached to hydrogen. For example, the strengths of these compounds decrease as H<sub>2</sub>Te > H<sub>2</sub>Se > H<sub>2</sub>S > H<sub>2</sub>O and HI > HBr > HCl >> HF. These trends can be explained by considering the effectiveness of the overlap of the hydrogen 1s orbital with the *p* orbital on the other atom. In the case of fluorine, the *p* orbital is a 2*p* orbital of small size. Therefore, the overlap of the orbitals is effective, and the H–F bond is a strong one. With H<sup>+</sup> removed, F<sup>-</sup> is comparable in strength as an electron pair donor to the water molecule and it has a negative charge. Accordingly, the proton donation by hydrogen fluoride (HF) is not extensive and the acid is weak. On the other hand, the 5*p* orbital on iodine does not overlap well with the 1s orbital on hydrogen so H<sup>+</sup> is easily lost to water making the ionization complete. Similar arguments apply to the H<sub>2</sub>X acids and predict correctly that H<sub>2</sub>Te is the most acidic of that series.

It is generally true that for a series of oxyacids of the same element (e.g., HClO<sub>4</sub>, HClO<sub>3</sub>, HClO<sub>2</sub>, and HOCl), the acid containing the element in the highest oxidation state will be the strongest acid. Thus, H<sub>2</sub>SO<sub>4</sub> is a stronger acid than H<sub>2</sub>SO<sub>3</sub>; HNO<sub>3</sub> is a stronger acid than HNO<sub>2</sub>; and HClO<sub>4</sub> is a stronger acid than HClO<sub>3</sub> which is in turn a stronger acid than HClO<sub>2</sub>. This behavior can be explained by considering the structure of molecules having the same central atom. In the case of H<sub>2</sub>SO<sub>4</sub> and H<sub>2</sub>SO<sub>3</sub>, the structures are



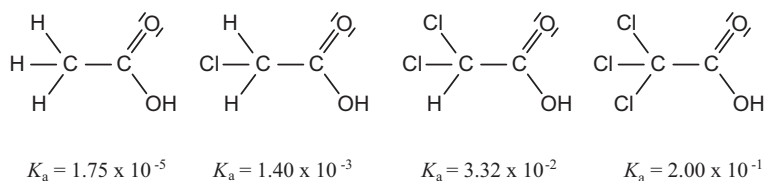
It is apparent that  $\text{H}_2\text{SO}_4$  has two oxygen atoms that are not bonded to hydrogen atoms whereas  $\text{H}_2\text{SO}_3$  has only one. Oxygen has the second highest electronegativity of any atom, and electrons are drawn toward those atoms by an *inductive effect*. The inductive effect caused by the two oxygen atoms having no hydrogen atoms attached removes electron density from the sulfur atom resulting in a slight positive charge, which is partially compensated by a general shift of electron density away from the O–H bonds.



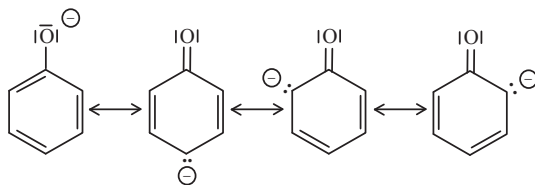
This shift in electron density makes the O–H bonds more polar and more susceptible to having a proton removed by a base. The same argument applies to the series  $\text{HClO}_4$ ,  $\text{HClO}_3$ ,  $\text{HClO}_2$ , and  $\text{HOCl}$ . Another way of viewing the inductive effect is to consider the stability of the anion after the proton is lost. Because there is some degree of double bonding to the oxygen atoms, there is greater delocalization of charge that stabilizes the anion after the proton is removed. This argument can also be used to explain the trend that as acids the strength varies as  $\text{H}_2\text{Te} > \text{H}_2\text{Se} > \text{H}_2\text{S}$  because the  $\text{HX}^-$  ions increase in size so the ease of accommodating the negative charge suggests this order.

When written as a general formula in which hydrogen atoms are bonded to oxygen atoms, the oxyacids can be represented as  $\text{XO}_a(\text{OH})_b$  where  $a = 0, 1, 2, \dots$  and  $b = 1, 2, 3, \dots$  depending on the acid. If  $a = 0$ , all of the oxygen atoms have a hydrogen atom attached, as in  $\text{H}_3\text{BO}_3$ , or more correctly,  $\text{B}(\text{OH})_3$ , and the acid is *very weak*. If  $a = 1$ , as in  $\text{H}_2\text{SO}_3$ ,  $\text{HNO}_2$ ,  $\text{HClO}_2$ , etc., there is one oxygen atom that does not have a hydrogen atom attached and the acid is *weak*. If  $a = 2$ , as in  $\text{H}_2\text{SO}_4$ ,  $\text{HNO}_3$ , and  $\text{HClO}_3$ , the acid is *strong*, at least in the first step of dissociation. If  $a = 3$ , as in  $\text{HClO}_4$ , there are three oxygen atoms that give an inductive effect and the acid is *very strong*. This line of reasoning correctly predicts that those acids for which  $a = 3$  will be stronger, as a group, than those for which  $a = 2$ ; those for which  $a = 2$  will be stronger than those for which  $a = 1$ ; etc.

The dissociation constants of the chloroacetic acids whose structures are shown below also illustrate the inductive effect produced by chlorine atoms.



Because Cl has a high electronegativity, it withdraws electron density from the O–H region of the molecule leaving the hydrogen atom with a slightly greater charge, making it easier to remove. In some cases, the loss of a proton is made more energetically favorable by the formation of an anion that has lower energy. For example, phenol,  $\text{C}_6\text{H}_5\text{OH}$ , is an acid for which  $K_a$  is  $1.1 \times 10^{-10}$ . On the other hand, ethanol,  $\text{C}_2\text{H}_5\text{OH}$  is not normally an acid ( $K_a \approx 10^{-17}$ ) except toward extremely strong bases such as  $\text{O}^{2-}$ ,  $\text{H}^-$ , or  $\text{NH}_2^-$ . Although both cases involve breaking an O–H bond, at least part of this difference can be attributed to the stability of the phenoxide ion for which several resonance structures can be drawn.



In general, the greater the number of contributing resonance structures, the more stable the species will be. This large number of resonance structures has the effect of producing an anion that has a lower energy than that of the alkoxide ion



for which resonance is not possible. Consequently, phenol is a stronger acid than the aliphatic alcohols.

For polyprotic acids such as  $\text{H}_3\text{PO}_4$  or  $\text{H}_3\text{AsO}_4$ , there is usually a factor of approximately  $10^5$  difference in successive  $K_a$  values. Phosphoric acid has dissociation constants that have the values  $K_{a1} = 7.5 \times 10^{-3}$ ,  $K_{a2} = 6.2 \times 10^{-8}$ , and  $K_{a3} = 1.0 \times 10^{-12}$ . The reason for this is that the first proton comes from a neutral molecule, the second from a  $-1$  ion, and the third from a  $-2$  ion. As a result of electrostatic attraction, it is energetically less favorable to remove  $\text{H}^+$  from species that are already negative. When considering the first and second ionization constants for the  $\text{H}_2\text{X}$  acids, it is seen that there is generally a factor of about  $10^7$ – $10^8$  difference in the values.

One of the general principles related to the Brønsted theory is that a stronger acid or base displaces or produces a weaker one. For example,  $\text{H}_2\text{SO}_4$  is a stronger acid than  $\text{HC}_2\text{H}_3\text{O}_2$  so the following reaction illustrates the principle.



In this reaction, the weaker acid  $\text{HC}_2\text{H}_3\text{O}_2$  has been produced by the stronger one,  $\text{H}_2\text{SO}_4$ . Similarly, a stronger base displaces or produces a weaker one. The following reaction illustrates this principle.



In this reaction, the stronger base,  $\text{OH}^-$ , displaces the weaker base,  $\text{NH}_3$ , from the proton by removing  $\text{H}^+$  from  $\text{NH}_4^+$ . Reactions involving these principles are not always simple as a result of differences in volatility of the compounds and the effects of solvation. For example, it is possible to produce  $\text{HCl}$  by heating a chloride with phosphoric acid even though  $\text{H}_3\text{PO}_4$  is a weaker acid than  $\text{HCl}$ . The reaction occurs as a result of the volatility of  $\text{HCl}(g)$ .



## 6.1.2 Factors Affecting Base Strength

If we consider bases as being the conjugates of acids, we can see some general trends about how base strength varies. For example,  $\text{H}_2\text{SO}_4$  is a strong acid and  $\text{H}_2\text{SO}_3$  is a weak acid. Therefore,  $\text{HSO}_4^-$  is a weaker base than is  $\text{HSO}_3^-$ . The stronger an acid is (more easily it donates protons), the weaker its conjugate will be as a base (weaker attraction for protons). Likewise,  $\text{SO}_3^{2-}$  will be a stronger base than  $\text{HSO}_3^-$  because  $\text{H}_2\text{SO}_3$  (the conjugate of  $\text{HSO}_3^-$ ) is a stronger acid than  $\text{HSO}_3^-$  (the conjugate acid of  $\text{SO}_3^{2-}$ ).

The strengths of bases can also be considered in terms of their charges. Consider, for example, the series of nitrogen bases  $\text{NH}_3 < \text{NH}_2^- < \text{NH}^{2-} < \text{N}^{3-}$ , which are written in the order of increasing base strength. Because the ideas about strengths of conjugates can be used (see above), it is very simple to see that a species having a  $-3$  charge will attract  $\text{H}^+$  more strongly than one having a  $-2$  charge. A similar argument concerning  $\text{PO}_4^{3-}$  and  $\text{SO}_4^{2-}$  indicates that the phosphate is much more basic than the sulfate ion.

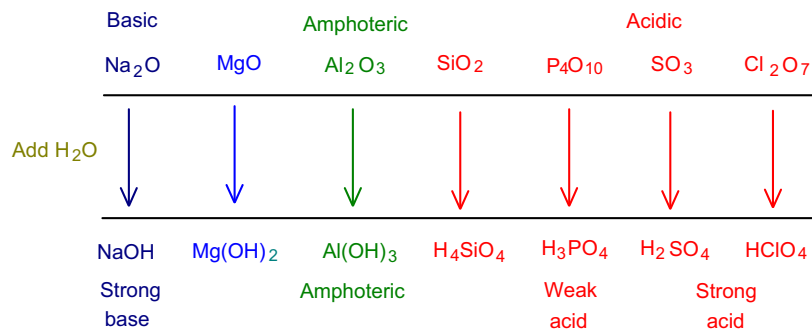
If we consider the  $\text{S}^{2-}$  and  $\text{O}^{2-}$  ions, we see that the charges are identical. However,  $\text{S}^{2-}$  is larger (190 pm) than  $\text{O}^{2-}$  (145 pm) so it will not interact as well with a small positively charged species such as  $\text{H}^+$ . As a result,  $\text{O}^{2-}$  is a stronger base than  $\text{S}^{2-}$  and  $\text{OH}^-$  is a stronger base than  $\text{HS}^-$  for the same reason. After all,  $\text{H}_2\text{S}$  is more acidic than  $\text{H}_2\text{O}$  so the conjugate  $\text{HS}^-$  would be a weaker base than  $\text{OH}^-$ .

Finally, if we consider  $\text{NH}_3$  and  $\text{PH}_3$  both of which are neutral species, we can still use the concept of size although this will be treated much more fully in a later section. In functioning as a base, the species provides a pair of electrons in an orbital where  $\text{H}^+$  becomes attached. Because  $\text{H}^+$  is a very small region of positive charge, it is attracted more strongly to a pair of electrons that are restricted to a small region of space. An electron pair is localized in a smaller orbital in the  $\text{NH}_3$  molecule than it is in  $\text{PH}_3$ . Accordingly,  $\text{NH}_3$  is a stronger base toward  $\text{H}^+$  than is  $\text{PH}_3$ .

## 6.1.3 Acids and Bases from Oxides

It was recognized very long ago that binary oxides give solutions that differ greatly when they are placed in water. For example, the oxides of nonmetals such as sulfur, halogens, phosphorus, etc., yield acidic solutions. Typical reactions can be shown as follows.





**FIGURE 6.1** The acid–base character of oxides.

However, the acid or base strength of the resulting solution depends on the nature of the oxide. In the examples above, the acids produced are strong acids, but the oxides of other nonmetals can yield weakly acidic solutions. Oxides that exhibit that behavior are those of carbon and phosphorus.



As a result of this acidic behavior of nonmetal oxides, such compounds are sometimes referred to as *acid anhydrides* (acids without water).

When they are placed in water, metal oxides produce basic solutions owing to the strongly basic nature of the oxide ion,



The reaction of  $\text{CaO}$  with water can be shown as

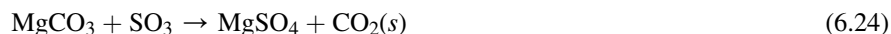


The acid–base character of oxides can be summarized as shown in [Figure 6.1](#).

Earlier we saw that water is not necessary for the transfer of a proton from gaseous  $\text{HCl}$  to gaseous  $\text{NH}_3$ . In a somewhat similar manner, an acidic oxide can react with a basic oxide to produce a salt. Examples of this type of reaction are the following.



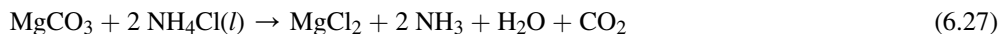
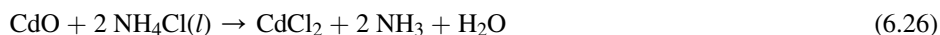
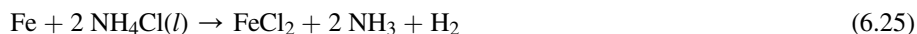
As a result of the difference in strength, replacement of a weaker acidic oxide by a stronger one is also possible. Such a reaction is



in which the weaker acidic oxide,  $\text{CO}_2$ , is displaced by the stronger,  $\text{SO}_3$ .

### 6.1.4 Molten Salt Protonic Acids

The ability of compounds to function as Brønsted acids or bases is not limited to aqueous solutions, and the reaction of gaseous  $\text{HCl}$  and  $\text{NH}_3$  has already been described. A great many compounds also behave as Brønsted acids and bases in the molten state. For example, molten  $\text{NH}_4\text{Cl}$  and pyridinium chloride (also known as pyridine hydrochloride,  $\text{C}_5\text{H}_5\text{NH}^+\text{Cl}^-$ ), readily undergo reactions that are typical of acids. In some of the early studies on this type of chemical behavior, a large number of reactions of molten  $\text{NH}_4\text{Cl}$  with metals and metal compounds were carried out. Typical reactions are illustrated in the following equations.



These reactions are proton transfer reactions that are typical of acids. For comparison, the equivalent reactions for gaseous HCl can be written as follows.

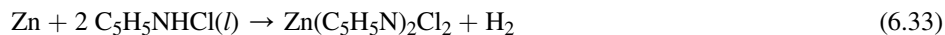


Toward certain metals, molten  $\text{NH}_4\text{Cl}$  is more reactive than gaseous HCl. The acidic properties of  $\text{NH}_4\text{Cl}$  are responsible for its use as a soldering flux (known as *sal ammoniac*) to remove oxide coatings.

Pyridine hydrochloride, which melts at 143–144 °C, has also been used as an acid in the molten state. Its acidity is illustrated by the following reactions.



However, with metals that complex strongly with pyridine, the pyridine may form a complex with the metal and thus not be lost from the system.



This type of behavior is not limited to amine hydrochlorides, and amine hydrobromides and hydrothiocyanates undergo similar reactions. Some of the reactions of amine hydrochlorides and hydrothiocyanates are useful in the preparation of coordination compounds containing the amines. The important characteristic is that ammonia or a protonated amine is a conjugate acid of a weak base and is able, therefore, to react as an acid.

### 6.1.5 Lewis Theory

The Brønsted theory explains acid–base reactions in terms of proton transfer. However, gaseous ammonia reacts with gaseous  $\text{BCl}_3$  to produce a white solid compound,  $\text{H}_3\text{N}:\text{BCl}_3$ . In the case of the reaction of  $\text{HCl}(g)$  with  $\text{NH}_3(g)$ , the unshared pair of electrons on the nitrogen atom is used to form a coordinate bond (a bond in which both electrons come from one of the atoms instead of one from each) to a proton. In the case of the reaction with  $\text{BCl}_3(g)$ , the unshared pair of electrons on the nitrogen atom is used to form a coordinate bond to boron in the  $\text{BCl}_3$  molecule. Although they are quite different, these cases involve the formation of a coordinate bond with the nitrogen atom in the  $\text{NH}_3$  molecule being the electron pair donor. However, the reaction with HCl involves proton transfer whereas the reaction with  $\text{BCl}_3$  does not. This situation leads us to examine a definition of acid–base reactions that does not involve proton transfer. Such an acid–base theory was advanced by G. N. Lewis in 1923.

Lewis examined four characteristics of acid–base reactions that are to be accounted for by any theory.

1. Neutralization—acids and bases combine rapidly with the loss of the characteristics of both.
2. Reactions with indicators—acids and bases provide characteristic reactions with indicators that change colors during neutralization.
3. Displacement reactions—a stronger acid or base will displace a weaker acid or base.
4. Catalytic activity—acids and bases frequently function as catalysts.

According to the Lewis theory, the reaction

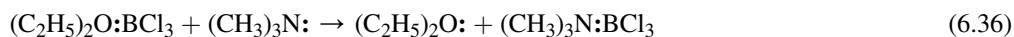


is a neutralization reaction because the electron donation character of  $\text{NH}_3$  is lost and the electron acceptance function of  $\text{BCl}_3$  is satisfied. In the *Lewis theory*, the *acid* and *base* are defined in terms of the behavior of *electrons* not protons. After all, bonding between species is described in terms of electrons, not protons. In the two cases described earlier, ammonia makes the unshared pairs of electrons available for the formation of coordinate bonds. It follows that a *base* is a substance that functions as an *electron pair donor*. An *acid* is a substance that functions as an *electron pair acceptor*. Although there are many known cases of indicators that change colors in the presence of Lewis acids or bases, these will not be described here. It is sufficient to point out that the titration of a Lewis acid with a Lewis base in an appropriate solvent (one that does not react with the acid or base) can be carried out using a suitable indicator. Many examples are described in the classic book *The Electronic Theory of Acids and Bases* by Luder and Zuffanti (see references).

We have seen examples of displacement reactions such as



in which the *weaker* base,  $\text{NH}_3$ , is displaced by the *stronger* one,  $\text{OH}^-$ . According to the Lewis theory, a reaction such as



can be considered as a displacement reaction in which the *stronger* base,  $(\text{CH}_3)_3\text{N}$ , displaces the *weaker* one,  $(\text{C}_2\text{H}_5)_2\text{O}$  from its compound. The  $\text{CO}_3^{2-}$  ion can be considered as an adduct formed from  $\text{CO}_2$  and  $\text{O}^{2-}$ . As a result, the reaction



represents the displacement of the weaker acidic oxide,  $\text{CO}_2$ , by the stronger acid,  $\text{SO}_3$ , which has a higher affinity for the  $\text{O}^{2-}$  ion. Numerous reactions illustrating this type of displacement will be shown in later sections.

If we consider the reaction



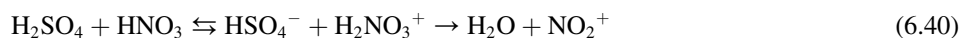
in which A and X are Lewis acids and B is a Lewis base, we see that an acid displacement takes place. This occurs because A is a stronger Lewis acid than X, and the weaker acid is displaced from its compound,  $\text{X}:\text{B}$ , by the stronger one. The same approach can be taken for base displacement reactions such as



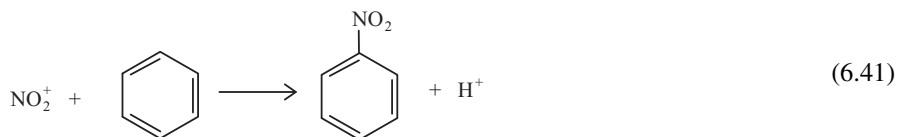
in which B and Y are Lewis bases and A is a Lewis acid. In this case, B is a stronger base than Y because B displaces Y from the compound  $\text{A}:\text{Y}$ .

A Lewis acid is a species that “seeks” a center of negative charge (considered to be an electron pair). Therefore, it is known as an *electrophile*, and Eq. (6.38) represents one electrophile displacing another in a reaction known as an *electrophilic substitution*. A Lewis base is a species that interacts with a center of positive charge (or deficiency of electrons) because of its unshared pair of electrons. Therefore, a Lewis base is known as a *nucleophile* and Eq. (6.39) represents a *nucleophilic substitution* reaction.

With regard to the characteristic of catalytic activity, we can consider the nitration of benzene, a reaction that is catalyzed by sulfuric acid. The first step in the process can be shown as



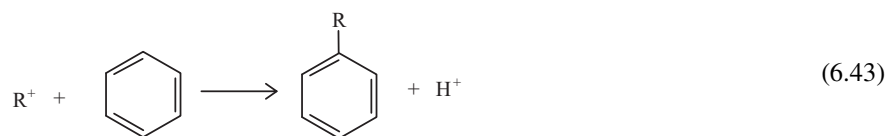
The nitronium ion,  $\text{NO}_2^+$ , is the species that attacks benzene.



The function of the acid catalyst is to generate a *positive* attacking species. In the *Friedel–Crafts reaction*,  $\text{AlCl}_3$  functions as an acid catalyst by increasing the concentration of the carbocation,  $\text{R}^+$ , which is the attacking species as shown in the following equations.







Although we will not show the equations here (the aldol condensation is such an example), the function of a base catalyst is to generate a *negative* attacking species.

Substances that are capable of functioning as Lewis bases include a wide variety of species that have one or more unshared pairs of electrons. They may be neutral molecules (e.g.,  $\text{NH}_3$ ,  $\text{H}_2\text{O}$ ,  $\text{R}_2\text{O}$ ,  $\text{ROH}$ ,  $\text{RSH}$ ,  $\text{H}_2\text{N}-\text{NH}_2$ , amines,  $\text{R}_3\text{P}$ ,  $\text{R}_3\text{PO}$ ,  $\text{SF}_4$ , etc.). Anions (e.g.,  $\text{OH}^-$ ,  $\text{H}^-$ ,  $\text{NH}_2^-$ ,  $\text{CN}^-$ ,  $\text{NO}_2^-$ ,  $\text{SCN}^-$ ,  $\text{C}_2\text{O}_4^{2-}$ , etc.) also have unshared pairs and function as Lewis bases.

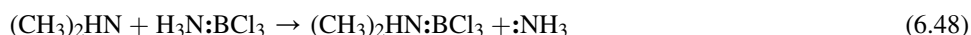
Substances that function as Lewis acids are those that are capable of accepting pairs of electrons. In  $\text{BCl}_3$ , the boron atom has only six electrons around it so it can accept an electron pair from a suitable electron pair donor. One class of Lewis acids includes molecules in which an atom does not have an octet of electrons around it (e.g.,  $\text{BCl}_3$ ,  $\text{BF}_3$ ,  $\text{AlCl}_3$ ,  $\text{BeCl}_2$ , etc.). Cations constitute another type of Lewis acid. Such ions as  $\text{H}^+$ ,  $\text{Fe}^{3+}$ ,  $\text{Ag}^+$ ,  $\text{Cr}^{3+}$ ,  $\text{Co}^{3+}$ ,  $\text{Pt}^{2+}$ ,  $\text{VO}^{2+}$ , and  $\text{BrF}_2^+$  are acidic because they can accept pairs of electrons.

Molecules in which the atoms have filled outer shells of electrons may also behave as Lewis acids if one of the atoms can expand its valence shell to hold more than eight electrons. This behavior is illustrated by  $\text{SbF}_5$  in the following reaction.



In this reaction,  $\text{SbF}_5$  accepts a pair of electrons from  $\text{F}^-$  even though it already has 10 electrons around it. This occurs because participation of another *d* orbital in bonding allows the valence shell to hold 12 electrons. This behavior does not occur for light elements because the central atom usually does not have *d* orbitals of low energy. Lewis acids of this type include molecules such as  $\text{BrF}_3$ ,  $\text{SF}_4$ ,  $\text{SnCl}_4$ ,  $\text{SiCl}_4$ , and  $\text{PCl}_5$ .

According to the Lewis acid–base theory, we can now consider reactions such as the following to be acid–base reactions.



Electron donation–acceptance reactions, which are considered to be Lewis acid–base interactions, also include the formation of coordination compounds, complex formation through hydrogen bonding, charge transfer complex formation, etc. It should be apparent that the Lewis theory of acids and bases encompasses a great deal of both inorganic and organic chemistry.

The products of acid–base interactions such as those shown in Eqs (6.45)–(6.48) are not properly considered as salts because they are not ionic compounds. Because in many cases these products are formed from two neutral molecules, they are more properly considered as addition compounds or *adducts* held together by the formation of coordinate covalent bonds. In that connection, they are similar to coordination compounds except that the latter ordinarily involve the formation of coordinate bonds to metal ions by the electron donors (*ligands*). There are some very useful generalizations that correlate to the stability of bonds during this type of acid–base interaction, and these are largely summarized by the *hard–soft acid–base* (HSAB) *principle*.

### 6.1.6 HSAB Principle

As a result of there being so many kinds of interactions that involve the donation and acceptance of electrons, the electronic theory of acids and bases pervades the whole of chemistry. In the 1950s, Ahrland, Chatt, and Davies had classified metals as Class A metals if they formed more stable complexes with the first element in a periodic group or Class B metals if they formed more stable complexes with the heavier elements in that group. Therefore, metals are classified as A or B based on which donor atom they prefer to bond to. A Class A metal ion such as  $\text{Cr}^{3+}$  interacts

preferentially with electrons donated by a nitrogen atom rather than a phosphorus atom. On the other hand, a Class B metal ion such as  $\text{Pt}^{2+}$  normally interacts more strongly with electrons donated by a phosphorus atom rather than a nitrogen atom. In these cases, the donor strength of the ligands is determined by the stability of the complexes. These ideas are summarized in the following table.

Donor Strength	
Class A metals	$\text{N} \gg \text{P} > \text{As} > \text{Sb} > \text{Bi}$
	$\text{O} \gg \text{S} > \text{Se} > \text{Te}$
	$\text{F} > \text{Cl} > \text{Br} > \text{I}$
Class B metals	$\text{N} \ll \text{P} > \text{As} > \text{Sb} > \text{Bi}$
	$\text{O} \ll \text{S} \approx \text{Se} \approx \text{Te}$
	$\text{F} < \text{Cl} < \text{Br} < \text{I}$

Thus,  $\text{Cr}^{3+}$  and  $\text{Co}^{3+}$  belong to Class A because they form more stable complexes with oxygen as the donor atom than with sulfur as the donor atom. On the other hand,  $\text{Ag}^+$  and  $\text{Pt}^{2+}$  belong to Class B because they form more stable complexes with P or S as the donor atom than with N or O as the donor atom.

The concept of hard and soft characteristics of acids and bases as discussed here was first put into systematic form by R. G. Pearson in the 1960s. According to the descriptions given by Pearson, soft bases are those electron pair donors that have high polarizability, low electronegativity, empty orbitals of low energy, or are easily oxidizable. Hard bases have the opposite properties. Soft acids are those having low positive charge, large size, and completely filled outer orbitals. Polarizability (the ability to distort the electron cloud of a species) depends upon these properties and low electronegativity. Hard acids have the opposite characteristics. Based on size and charge, we would expect that typical hard acids would be species such as  $\text{Cr}^{3+}$ ,  $\text{Co}^{3+}$ ,  $\text{Be}^{2+}$ ,  $\text{H}^+$ , and similar species. Soft acids would include species such as  $\text{Ag}^+$ ,  $\text{Hg}^{2+}$ ,  $\text{Pt}^{2+}$ , or uncharged metal atoms. Obviously, such a distinction is made on a qualitative basis, and the classification of some species may be somewhat uncertain. Tables 6.2 and 6.3 show some typical acids and bases classified according to their hard or soft character. (Lists are based on those in Pearson, R. G. *J. Chem. Educ.* **1968**, 45, 581.)

The general rule relating to the interaction between acids and bases is that *the most favorable interactions occur when the acid and base have similar electronic character*. Thus, hard acids interact *preferentially* with hard bases, and soft acids interact *preferentially* with soft bases. This can be explained in terms of type of bonding. The hard acids interact with hard

**TABLE 6.2** Lewis Bases

Hard	Borderline	Soft
$\text{OH}^-$ , $\text{H}_2\text{O}$ , $\text{F}^-$	$\text{C}_5\text{H}_5\text{N}$ , $\text{N}_3^-$ , $\text{Br}^-$ , $\text{NO}_2^-$	$\text{RS}^-$ , $\text{RSH}$ , $\text{R}_2\text{S}$
$\text{SO}_4^{2-}$ , $\text{Cl}^-$ , $\text{PO}_4^{3-}$		$\text{I}^-$ , $\text{SCN}^-$ , $\text{CN}^-$
$\text{ClO}_4^-$ , $\text{RO}^-$ , $\text{ROH}$ , $\text{R}_2\text{O}$		$\text{CO}$ , $\text{H}^-$ , $\text{R}^-$
$\text{NH}_3$ , $\text{RNH}_2$ , $\text{N}_2\text{H}_4$		$\text{R}_3\text{P}$ , $\text{R}_3\text{As}$ , $\text{C}_2\text{H}_4$

**TABLE 6.3** Lewis Acids

Hard	Borderline	Soft
$\text{H}^+$ , $\text{Li}^+$ , $\text{Na}^+$ , $\text{K}^+$	$\text{Fe}^{2+}$ , $\text{Co}^{2+}$ , $\text{Ni}^{2+}$ , $\text{Zn}^{2+}$	$\text{Cu}^+$ , $\text{Ag}^+$ , $\text{Au}^+$ , $\text{Ru}^+$
$\text{Be}^{2+}$ , $\text{Mg}^{2+}$ , $\text{Ca}^{2+}$ , $\text{Mn}^{2+}$	$\text{Cu}^{2+}$ , $\text{Sb}^{3+}$ , $\text{SO}_2$ , $\text{NO}^+$	$\text{Pd}^{2+}$ , $\text{Cd}^{2+}$ , $\text{Pt}^{2+}$ , $\text{Hg}^{2+}$
$\text{Al}^{3+}$ , $\text{Sc}^{3+}$ , $\text{La}^{3+}$ , $\text{Cr}^{3+}$		$\text{GaCl}_3$ , $\text{RS}^+$ , $\text{I}^+$ , $\text{Br}^+$
$\text{Co}^{3+}$ , $\text{Fe}^{3+}$ , $\text{Si}^{4+}$ , $\text{Tl}^{4+}$		$\text{O}$ , $\text{Cl}$ , $\text{Br}$ , $\text{I}$ , $\text{N}$
$\text{Be}(\text{CH}_3)_2$ , $\text{BF}_3$ , $\text{HCl}$		Uncharged metal atoms

bases primarily by ionic or polar interactions, which will be favored by high charge and small size of both the acid and base. Soft acids and soft bases interact primarily by forming covalent bonds that may involve distortion of electron clouds (polarization). Frequently, these interactions involve bonding between neutral molecules, and sharing of electron density is more favorable when the orbitals of the donor and acceptor atoms are of similar size and energy. This principle should not be interpreted to indicate that hard acids will *not* interact with soft bases, but only that *more favorable* interactions occur between acids and bases that have similar electronic character.

A modification of the HSAB approach was first explained by C. K. Jørgensen. We will consider an actual example to make this idea clear. The  $\text{Co}^{3+}$  ion is a hard Lewis acid. However, when  $\text{Co}^{3+}$  is bonded to five cyanide ions, a more stable complex results when the sixth group is iodide than when it is fluoride. In other words,  $[\text{Co}(\text{CN})_5\text{I}]^{3-}$  is stable whereas  $[\text{Co}(\text{CN})_5\text{F}]^{3-}$  is not. At first this seems like a contradiction that the soft  $\text{I}^-$  bonds more strongly to the “hard” acid,  $\text{Co}^{3+}$ . However, the five  $\text{CN}^-$  ions have made the  $\text{Co}^{3+}$  in the complex much softer than an isolated  $\text{Co}^{3+}$  ion. Thus, when five  $\text{CN}^-$  ions are attached, the cobalt ion behaves as a soft acid. This effect is known as the *symbiotic effect*, and it indicates that whether a species appears to be hard or soft depends on the other groups attached and their character.

## 6.1.7 Applications of the Hard–Soft Interaction Principle

The HSAB principle is not restricted to the usual types of Lewis acid–base interactions. It is a guiding principle for all types of interactions that species of similar electronic character interact best. Accordingly, we will refer to the principle as the hard–soft interaction principle (HSIP). We have already seen some applications of this principle (such as the relative strength of HF and HI), but we now consider a number of other types of applications.

### 6.1.7.1 Hydrogen Bonding

The HSIP can be applied in a qualitative way to hydrogen bonding interactions. Because the hydrogen end (a proton) of an O–H bond is hard, stronger hydrogen bonds should be formed when the electron donor atom is a hard Lewis base. The occurrence of hydrogen bonding in the hydrogen compounds of the first long group of elements has already been cited as being responsible for their high boiling points (see Chapter 4). Hydrogen bonding is much more extensive in liquid  $\text{NH}_3$ ,  $\text{H}_2\text{O}$ , and HF than it is in  $\text{PH}_3$ ,  $\text{H}_2\text{S}$ , and HCl. Unshared pairs of electrons on second row atoms are contained in larger orbitals and do not interact as well with the very small H nucleus. As a result, liquid  $\text{NH}_3$ ,  $\text{H}_2\text{O}$ , and HF have higher boiling points than do liquid  $\text{PH}_3$ ,  $\text{H}_2\text{S}$ , and HCl.

Another interesting application of the HSIP is afforded by considering the interaction of an alcohol with acetonitrile,  $\text{CH}_3\text{CN}$ , and trimethylamine,  $(\text{CH}_3)_3\text{N}$ , each of which has an unshared pair of electrons on a nitrogen atom. The dipole moments of these molecules are 3.44 and 0.7 D, respectively. However, nitriles are soft bases and amines are hard bases. In this case, when  $\text{CH}_3\text{OH}$  is hydrogen bonded to  $\text{CH}_3\text{CN}$  and  $(\text{CH}_3)_3\text{N}$ , the bonds have energies of about  $6.3 \text{ kJ mol}^{-1}$  and  $30.5 \text{ kJ mol}^{-1}$ , respectively, in accord with the predictions of HSIP.

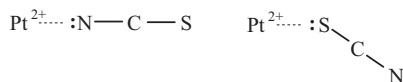
When phenol,  $\text{C}_6\text{H}_5\text{OH}$ , forms hydrogen bonds to  $(\text{C}_2\text{H}_5)_2\text{O}$  in dilute solution in a nonpolar solvent such as  $\text{CCl}_4$ , the OH band in the infrared spectrum is shifted by about  $280 \text{ cm}^{-1}$  from where it is in gaseous phenol, and the hydrogen bonds have energies of about  $22.6 \text{ kJ mol}^{-1}$ . When phenol hydrogen bonds to  $(\text{C}_2\text{H}_5)_2\text{S}$ , the corresponding values for these parameters are about  $250 \text{ cm}^{-1}$  and  $15.1 \text{ kJ mol}^{-1}$ . Ethers are considered to be hard bases whereas alkyl sulfides are soft.

If our qualitative predictions regarding softness of the base are valid, we should expect that  $\text{C}_6\text{H}_5\text{OH}$  hydrogen bonded to the bases  $(\text{C}_6\text{H}_5)_3\text{P}$  and  $(\text{C}_6\text{H}_5)_3\text{As}$  would also follow the trend indicated above. In fact, when the OH group of phenol is hydrogen bonded to  $(\text{C}_6\text{H}_5)_3\text{P}$ , the OH stretching band is shifted by  $430 \text{ cm}^{-1}$ , and when it is hydrogen bonded to  $(\text{C}_6\text{H}_5)_3\text{As}$  it is shifted by  $360 \text{ cm}^{-1}$ . This is exactly as expected because arsenic is a softer electron pair donor than phosphorus. When phenol forms hydrogen bonds to  $\text{CH}_3\text{SCN}$ , the hydrogen bond energy is  $15.9 \text{ kJ mol}^{-1}$  and the OH band is shifted by  $146 \text{ cm}^{-1}$ . However, when phenol forms hydrogen bonds to  $\text{CH}_3\text{NCS}$ , the OH stretching band is shifted by  $107 \text{ cm}^{-1}$  and the hydrogen bond energy is  $7.1 \text{ kJ mol}^{-1}$ . The sulfur end of the SCN group is a soft electron donor and the nitrogen end is significantly harder. The energies of the hydrogen bonds and the spectral shifts reflect this difference.

### 6.1.7.2 Linkage Isomers

Ions such as  $\text{SCN}^-$  have two potential electron donor atoms. When bonding to metal ions, the bonding mode may be predicted by means of the HSIP. For example, when  $\text{SCN}^-$  bonds to  $\text{Pt}^{2+}$ , it bonds through the sulfur atom, but when it bonds to  $\text{Cr}^{3+}$ , it bonds through the nitrogen atom. This is in accord with the Class A and Class B behavior of these metal

ions described earlier. However, in some cases steric effects can cause a change in bonding mode in  $\text{Pt}^{2+}$  complexes. When bonding is through the nitrogen atom, the arrangement is linear, but it is bent when bonding is through the sulfur atom.



When three very large groups such as  $\text{As}(\text{C}_6\text{H}_5)_3$  are bound in the other three positions around  $\text{Pt}^{2+}$ , the steric effects can cause  $\text{SCN}^-$  to bond through the nitrogen atom.

### 6.1.7.3 Solubility

One of the simplest applications of the HSIP principle is related to solubility. The rule, “like dissolves like,” is a manifestation that solute particles interact best with solvent molecules that have similar characteristics. Small, highly charged particles or polar molecules are solvated best by solvents containing small, highly polar molecules. Large solute particles having low polarity are solvated best by solvent molecules having similar characteristics. Consequently,  $\text{NaCl}$  is soluble in water whereas sulfur,  $\text{S}_8$ , is not. On the other hand,  $\text{NaCl}$  is insoluble in  $\text{CS}_2$  but  $\text{S}_8$  dissolves in  $\text{CS}_2$ .  $\text{NaCl}$  dissolves to the extent of 35.9 g in 100 g of water at  $25^\circ\text{C}$ . The solubilities of  $\text{NaCl}$  in  $\text{CH}_3\text{OH}$ ,  $\text{C}_2\text{H}_5\text{OH}$ , and *iso*- $\text{C}_3\text{H}_7\text{OH}$  are 0.237, 0.0675, and 0.0041 g in 100 g of solvent, respectively, at  $25^\circ\text{C}$ . In the case of the alcohols, as the chain length increases, the organic part of the molecule dominates the OH functional group so that these compounds become progressively poorer solvents for ionic solutes. Solubility is a complicated issue because in some cases the energy necessary to create a cavity in the solvent is a dominant factor. For example, because water has a high cohesion energy, it is difficult to form a cavity in the solvent in which a molecule such as  $\text{CH}_4$  can reside. The interactions between water molecules are much stronger than those between water and  $\text{CH}_4$ .

The HSIP also applies to precipitation of solids from solution. Ionic solids precipitate best from aqueous solutions when the ions are of similar size, preferably with the two ions having charges of the same magnitude. For the reaction

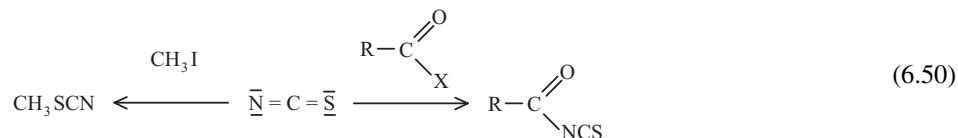


the enthalpy changes are 0, 66.9, 58.6, and  $-20.9 \text{ kJ mol}^{-1}$  for  $\text{LiF}$ ,  $\text{LiI}$ ,  $\text{CsF}$ , and  $\text{CsI}$ , respectively. These data indicate that if a solution contains  $\text{Li}^+$ ,  $\text{Cs}^+$ ,  $\text{F}^-$ , and  $\text{I}^-$  the precipitation of  $\text{LiF}$  and  $\text{CsI}$  is more favorable than that of  $\text{LiI}$  and  $\text{CsF}$ . We can conclude that the small, hard  $\text{Li}^+$  ion interacts better with  $\text{F}^-$  and the large, soft  $\text{Cs}^+$  interacts better with the large  $\text{I}^-$ .

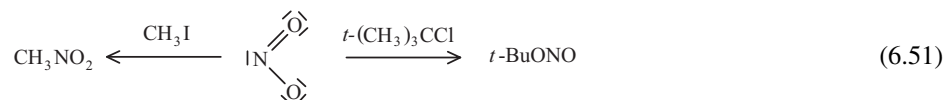
The application of the HSIP principle is of considerable importance in preparative coordination chemistry because some solid complexes are stable only when they are precipitated using a counter ion conforming to the above rule. For example,  $[\text{CuCl}_5]^{3-}$  is not stable in aqueous solution but can be isolated as solid  $[\text{Cr}(\text{NH}_3)_6][\text{CuCl}_5]$ . Attempts to isolate solid compounds containing the complex ion  $[\text{Ni}(\text{CN})_5]^{3-}$  as  $\text{K}_3[\text{Ni}(\text{CN})_5]$  yield  $\text{KCN}$  and  $\text{K}_2[\text{Ni}(\text{CN})_4]$  instead. It has been found, however, that when counter ions such as  $[\text{Cr}(\text{NH}_3)_6]^{3+}$  or  $[\text{Cr}(\text{en})_3]^{3+}$  are used, solids containing the  $[\text{Ni}(\text{CN})_5]^{3-}$  ion are obtained.

### 6.1.7.4 Reactive Site Preference

We have already used the HSIP principle as it applies to linkage isomers in metal complexes. This application to bonding site preference can also be used to predict the products of some reactions. For example, the reactions of organic compounds also obey the principle when reacting with nucleophiles such as  $\text{SCN}^-$  or  $\text{NO}_2^-$ . Consider the reactions illustrated in Eq. (6.50).



In this case, the acyl group in the acidic species  $\text{RCOX}$  is a hard acid and reacts with the nitrogen end of  $\text{SCN}^-$  to form an acyl isothiocyanate. The soft methyl group in methyl iodide bonds to the S atom and forms methyl thiocyanate. Consider the following reactions of  $\text{NO}_2^-$ .



Here, the  $t\text{-(CH}_3)_3\text{C}^+$  carbocation is a hard Lewis acid so the product is determined by its interaction with the oxygen (harder) electron donor in  $\text{NO}_2^-$ . In the reaction with  $\text{CH}_3\text{I}$ , the product is nitromethane, showing the softer character of the methyl group.

Another reaction that illustrates the HSIP is the reaction of  $\text{PCl}_3$  with  $\text{AsF}_3$ .



Although both arsenic and phosphorus are soft, arsenic is softer. Likewise, Cl is softer than F, so we predict that this reaction will take place as written, and this is verified experimentally. If we consider the reaction between  $\text{KSiH}_3$  and  $\text{CH}_3\text{Cl}$ , we might propose the following possibilities.



However,  $\text{K}^+$  and  $\text{Cl}^-$  are hard, and the  $\text{CH}_3$  and  $\text{SiH}_3$  groups are considerably softer. Thus, Eq. (6.55) represents the actual reaction between  $\text{KSiH}_3$  and  $\text{CH}_3\text{Cl}$ . Using the HSIP we can predict that the following reactions would occur as written.



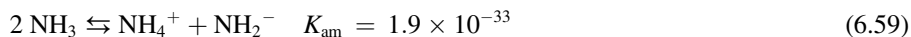
From these examples, we see that the HSIP is useful in many areas of chemistry.

## 6.2 NONAQUEOUS SOLVENTS

Although water is the most common solvent used in inorganic chemistry, other solvents are useful for a variety of reasons. For example, it is not possible to work with a base stronger than  $\text{OH}^-$  in aqueous solutions because a stronger base will react with water to produce  $\text{OH}^-$ . In liquid ammonia, it is possible to utilize the amide ion,  $\text{NH}_2^-$  as a base that is stronger than  $\text{OH}^-$ . A similar situation exists with acids because the strongest acid that can exist in aqueous solutions is  $\text{H}_3\text{O}^+$ . Because of these differences, it is often possible to carry out reactions in a nonaqueous solvent that would be impossible in water. However, work using nonaqueous solvents is frequently less convenient than aqueous solution chemistry. Several of the commonly used nonaqueous solvents are gases at room temperature (e.g.,  $\text{NH}_3$ ,  $\text{SO}_2$ , etc.), some are highly toxic (e.g.,  $\text{HCN}$  and  $\text{H}_2\text{S}$ ), and some react easily with traces of moisture (e.g.,  $\text{SOCl}_2$ ,  $\text{H}_2\text{SO}_4$ , and acetic anhydride,  $(\text{CH}_3\text{CO})_2\text{O}$ ). As a result, special techniques are often required in nonaqueous solvent chemistry. The physical properties of some of the commonly used nonaqueous solvents are shown in Table 6.4, and solubilities of inorganic compounds in some solvents have been described in Chapter 5.

### 6.2.1 The Solvent Concept

The fact that water undergoes some autoionization suggests that perhaps other solvents behave in a similar way. For predicting the products of reactions, it may not be important whether such ionization *actually* takes place. This will be discussed more fully in a later section. The autoionization of liquid ammonia can be written as



Just as the cation produced by dissociation of water ( $\text{H}_3\text{O}^+$ ) is the acidic species in aqueous solutions, the  $\text{NH}_4^+$  ion is the acidic species in liquid ammonia. Similarly, the amide ion,  $\text{NH}_2^-$ , is the base in liquid ammonia just as  $\text{OH}^-$  is the basic species in water. Generalization to other nonaqueous solvents leads to the *solvent concept* of acid–base behavior. It can be stated simply as: *A substance that increases the concentration of the cation characteristic of the solvent is an acid and a substance that increases the concentration of the anion characteristic of the solvent is a base.* Consequently,  $\text{NH}_4\text{Cl}$  is an acid in liquid ammonia and  $\text{NaNH}_2$  is a base in that solvent. Neutralization becomes the reaction of the cation and

**TABLE 6.4** Properties of Some Nonaqueous Solvents

Solvent	m.p., °C	b.p., °C	Dipole Moment, D <sup>a</sup>	Dielectric Constant
HCN	-13.4	25.7	2.8	114.9
H <sub>2</sub> SO <sub>4</sub>	10.4	338	—	100
HF	-83	19.4	1.9	83.6
H <sub>2</sub> O	0.0	100	1.85	78.5
N <sub>2</sub> H <sub>4</sub>	2.0	113.5	1.83	51.7
CH <sub>3</sub> OH	-97.8	65.0	1.68	31.5
NH <sub>3</sub>	-77.7	-33.4	1.47	26.7
(CH <sub>3</sub> CO) <sub>2</sub> O	-71.3	136.4	—	20.5
SO <sub>2</sub>	-75.5	-10.0	1.61	15.6
H <sub>2</sub> S	-85.5	-60.7	1.10	10.2
HSO <sub>3</sub> F	-89	163	—	—

<sup>a</sup>1 D = 3.336 × 10<sup>-30</sup> C m = 10<sup>-18</sup> esu cm.

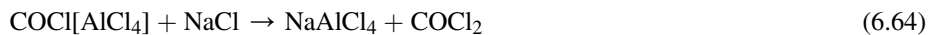
anion characteristic of the particular solvent to produce unionized solvent. For example, in liquid ammonia the following is a neutralization.



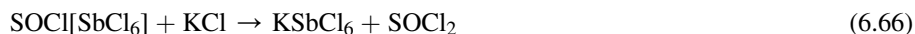
The solvent concept for nonaqueous solvents works exactly like the Arrhenius theory does for aqueous solutions. Autoionization and typical neutralization reactions can be shown as follows for several solvents. For liquid SO<sub>2</sub>,



For liquid phosgene, COCl<sub>2</sub>,



For liquid SOCl<sub>2</sub>,



For liquid N<sub>2</sub>O<sub>4</sub>,



*Amphoterism*, the ability to react as both an acid and a base, is also exhibited by substances in nonaqueous solvents. For example, in aqueous solutions Zn<sup>2+</sup> behaves as shown in the following equations.



When a base (such as NaOH) is added to Zn(OH)<sub>2</sub>, the reaction is



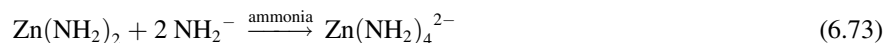
When an acid is added to  $\text{Zn}(\text{OH})_2$ , the reaction is



In liquid ammonia,  $\text{Zn}^{2+}$  also produces a precipitate with the  $\text{NH}_2^-$  anion.



With the addition of additional base containing  $\text{NH}_2^-$ ,



The addition of an acid ( $\text{NH}_4^+$ ) results in the reaction



It is readily apparent that  $\text{Zn}(\text{OH})_2$  is amphoteric in aqueous solution and that  $\text{Zn}(\text{NH}_2)_2$  behaves in an analogous way in liquid ammonia. Other examples of amphoterism in nonaqueous solvents are  $\text{Al}_2(\text{SO}_3)_3$  in liquid  $\text{SO}_2$  and  $\text{Zn}(\text{C}_2\text{H}_3\text{O}_2)_2$  in glacial acetic acid.

## 6.2.2 The Coordination Model

According to the autoionization patterns presented above, it might be supposed that  $\text{OPCl}_3$  could ionize slightly as represented by the equation



Because the addition of  $\text{FeCl}_3$  to liquid  $\text{OPCl}_3$  increases the concentration of the  $\text{OPCl}_2^+$  cation, ferric chloride is an acid in liquid  $\text{OPCl}_3$ , which can be shown according to the solvent concept as

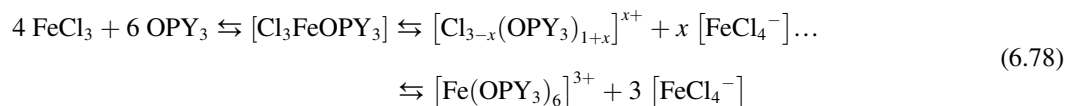


The simplest way to explain this behavior is to assume that some  $\text{Cl}^-$  is present because of autoionization or that removal of that  $\text{Cl}^-$  by complexing with  $\text{FeCl}_3$  causes the system represented by Eq. (6.75) to be shifted to the right causing more  $\text{OPCl}_2^+$  to be formed.

It is known, however, that toward some Lewis acids the oxygen atom in  $\text{OPCl}_3$  is a better electron pair donor (more basic site) than are the chlorine atoms. For example, when  $\text{OPCl}_3$  forms a complex with  $\text{AlCl}_3$  the bonding is through the oxygen atom rather than through one of the chlorine atoms. Moreover, when  $\text{AlCl}_3$  containing  $^{36}\text{Cl}$  is used, there is no isotope exchange with the chlorine atoms in  $\text{OPCl}_3$ , which would occur if  $\text{OPCl}_3$  undergoes autoionization. In spite of there being no chloride exchange, the presence of  $\text{AlCl}_4^-$  can be demonstrated. It is the result of a substitution reaction that can be represented as



According to the solvent concept, it might be presumed that some complex such as  $\text{Cl}_3\text{Fe}-\text{ClOPCl}_2$  would be formed when  $\text{OPCl}_3$  reacts with  $\text{FeCl}_3$ . However, this is inconsistent with the fact that the oxygen atom is a better electron donor than the chlorine atoms. Recognizing this problem, R. S. Drago and coworkers studied the interaction of  $\text{FeCl}_3$  with triethyl phosphate,  $\text{OP}(\text{OC}_2\text{H}_5)_3$ . Clearly, there is no possibility of electron donation by chlorine in this system because no chlorine atoms are present in triethyl phosphate. Any  $\text{FeCl}_4^-$  formed must result from dissociation of  $\text{FeCl}_3$  rather than from chloride ions formed by autoionization of the  $\text{OP}(\text{OC}_2\text{H}_5)_3$ . Spectral studies showed clearly that  $\text{FeCl}_4^-$  was present in solutions of  $\text{FeCl}_3$  in  $\text{OP}(\text{OC}_2\text{H}_5)_3$ . Furthermore, the spectra were similar for solutions of  $\text{FeCl}_3$  in both  $\text{OPCl}_3$  and  $\text{OP}(\text{OC}_2\text{H}_5)_3$ . These results indicate that some dissociation of  $\text{FeCl}_3$  must occur to produce  $\text{Cl}^-$  and cationic species containing iron(III). Such cations may have one or more chloride ions in  $\text{FeCl}_3$  replaced by neutral solvent molecules. It was suggested that the correct representation of the interaction between  $\text{FeCl}_3$  and  $\text{OPCl}_3$  or  $\text{OP}(\text{OC}_2\text{H}_5)_3$  (both represented as  $\text{OPY}_3$ ) is



These equations show that  $\text{FeCl}_4^-$  can form by *coordination* of the solvent rather than by postulating solvent autoionization according to the solvent concept. In the series of reactions represented by Eq. (6.78), nucleophilic substitution occurs

in which a solvent molecule replaces a chloride ion that subsequently interacts with  $\text{FeCl}_3$ . Undoubtedly, a similar situation exists for other reactions in which autoionization *appears* to occur. Autoionization probably occurs only in solvents in which a proton that is strongly solvated is transferred (e.g.,  $\text{H}_2\text{O}$ ,  $\text{HF}$ ,  $\text{NH}_3$ , etc.). Although the solvent concept is useful in a *formal* way, it is unlikely that autoionization occurs for a solvent such as liquid  $\text{SO}_2$ . However, many reactions take place to give the products that would be predicted *if* autoionization had occurred. We will now describe the chemistry of three of the most extensively studied nonaqueous solvents.

### 6.2.3 Liquid Ammonia

Because of the similarity of liquid ammonia to water, there has been an extensive chemistry of liquid ammonia for many years. Although  $\text{NH}_3$  boils at  $-33.4^\circ\text{C}$ , it has a rather high heat of vaporization because of intermolecular hydrogen bonding. As a result, the liquid does not evaporate very readily and that permits many reactions to be carried out at room temperature in a Dewar flask. Alternatively, reactions may also be carried out under pressure.

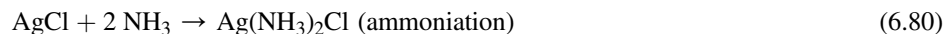
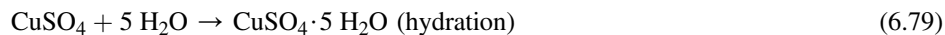
As a general rule, organic compounds are usually more soluble in liquid ammonia than they are in water. Inorganic salts are usually more soluble in water unless the cation forms stable complexes with  $\text{NH}_3$ . For example,  $\text{AgCl}$  is more soluble in liquid ammonia than it is in water because of the stability of  $\text{Ag}(\text{NH}_3)_2^+$ . Because the basic species in liquid  $\text{NH}_3$  is  $\text{NH}_2^-$ , reactions that involve strongly basic materials can frequently be carried out in liquid  $\text{NH}_3$  more readily than they can in water because  $\text{NH}_2^-$  is a stronger base than  $\text{OH}^-$ . Some of the important physical properties of liquid  $\text{NH}_3$  are shown in Table 6.5.

### 6.2.4 Reactions in Liquid Ammonia

Although ammonia itself is basic, it acts as an acid toward the strongest bases such as  $\text{H}^-$ ,  $\text{O}^{2-}$ , or  $\text{N}^{3-}$ . As we describe some of the reactions of liquid ammonia, the similarity to corresponding reactions of water should become apparent.

#### 6.2.4.1 Ammoniation Reactions

When solids crystallize from aqueous solutions, many are obtained as hydrates. In an analogous way, ammonia forms ammoniates.



In some cases, the  $\text{NH}_3$  is coordinated to a metal (as above) and in others it is present as ammonia of crystallization as is the case with  $\text{NBr}_3 \cdot 6 \text{NH}_3$ .

#### 6.2.4.2 Ammonolysis Reactions

A broad classification of reactions is *solvolysis*. In hydrolysis reactions, water molecules are split or *lysed* as illustrated by the reaction

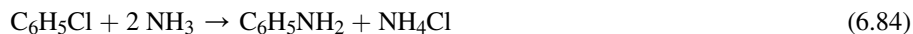


**TABLE 6.5** Physical Properties of Liquid  $\text{NH}_3$

Melting point	$-77.7^\circ\text{C}$
Boiling point	$-33.4^\circ\text{C}$
Density at $-33.4^\circ\text{C}$	$0.683 \text{ g cm}^{-3}$
Heat of fusion	$5.98 \text{ kJ mol}^{-1}$
Heat of vaporization	$22.84 \text{ kJ mol}^{-1}$
Dipole moment	1.47 D
Dielectric constant	22
Specific conductance at $-35^\circ\text{C}$	$2.94 \times 10^{-7} \text{ ohm}^{-1}$



The following equations illustrate ammonolysis reactions:



### 6.2.4.3 Acid–Base Reactions

According to the solvent concept, the acidic species characteristic of liquid ammonia is  $\text{NH}_4^+$  and the basic species is  $\text{NH}_2^-$ . Neutralization reactions in liquid ammonia thus become equivalent to the reaction of these ions.



Such a process occurs when  $\text{NH}_4\text{Cl}$  reacts with  $\text{KNH}_2$  in liquid ammonia.

Because liquid  $\text{NH}_3$  is a base, the ionization with even weak acids such as acetic acid goes to completion in that solvent.

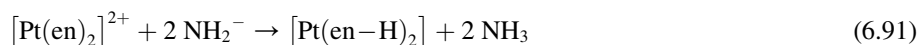


As a result, acetic acid is a strong acid in liquid ammonia owing to the basic character of the solvent. Urea,  $\text{OC}(\text{NH}_2)_2$ , is a weak base in water, but it behaves as an acid in liquid  $\text{NH}_3$ . The amide ion,  $\text{NH}_2^-$ , is the basic species in liquid ammonia. Although it is a stronger base than  $\text{OH}^-$ , it is weaker than  $\text{H}^-$ ,  $\text{O}^{2-}$ ,  $\text{NH}^{2-}$ , etc., as illustrated by the following equations:

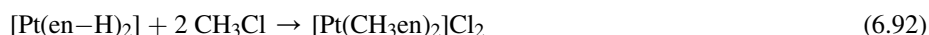


### 6.2.4.4 Deprotonation Reactions

Ethylenediamine,  $\text{H}_2\text{NCH}_2\text{CH}_2\text{NH}_2$ , (abbreviated as en in writing formulas for complexes) forms many stable complexes with metal ions. One such complex is  $[\text{Pt}(\text{en})_2]^{2+}$ . The amide ion will remove protons from the coordinated en so in liquid ammonia the reaction



occurs (where en–H represents an en molecule from which a proton has been removed). The deprotonated species undergoes a reaction with methyl chloride,



By taking advantage of the basicity of  $\text{NH}_2^-$  in liquid  $\text{NH}_3$ , it has been possible to carry out reactions on coordinated ligands to prepare complexes that are not obtainable by other means in aqueous solutions.

### 6.2.4.5 Solutions of Metals in Ammonia

One of the striking differences between water and ammonia is the behavior of alkali metals toward these solvents. With water, these metals react rapidly to liberate hydrogen.



However, liquid ammonia will dissolve these metals without reaction to produce solutions having many unusual properties. As shown in [Table 6.6](#), Group IA metals are quite soluble in liquid ammonia.

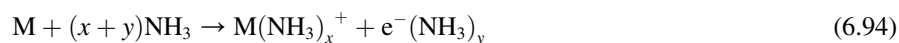
Evaporation of solutions of these metals in  $\text{NH}_3$  leads to recovery of the metal, usually as a solvated species such as  $\text{M}(\text{NH}_3)_6$ . Great care must be taken to assure that the ammonia is pure because these metals will react with traces of water or acidic species.

**TABLE 6.6** Solubility of Alkali Metals in Ammonia

Metal	Temperature, °C	Molality of Saturated Solution	Temperature, °C	Molality of Saturated Solution
Li	0	16.31	-33.2	15.66
Na	0	10.00	-33.5	10.93
K	0	12.4	-33.2	11.86
Cs	0	—	-50.0	25.1

The dilute solutions of alkali metals in liquid  $\text{NH}_3$  are less dense than the pure solvent, indicating that some expansion of the solvent has taken place as the metal dissolves. Electrical conductivities of the dilute solutions (which are blue in color) are characteristic of 1:1 electrolytes that are completely dissociated. At concentrations greater than about 1 M, the solutions are bronze in color, and they have conductivities as high as metals. The solutions are paramagnetic although the magnetic susceptibility decreases with increasing concentration, indicating that the *fraction* of free electrons is higher in more dilute solutions.

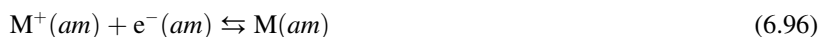
Hundreds of studies on these solutions have been carried out using electrical conductivity, magnetic susceptibility, NMR, volume expansion, spectroscopy (visible and infrared), and other techniques. The results of these studies indicate that in dilute solutions the metal dissociates to produce solvated metal ions and solvated electrons as illustrated by the equation



which can also be represented as



The solvated electrons are rather loosely bound to  $\text{NH}_3$  molecules, and they form cavities in the solvent. These structures are responsible for the volume expansion that occurs when the metals are dissolved, and they lead to solutions that are less dense than the pure solvent. In order to account for the decrease in paramagnetism at higher concentrations, it is proposed that some pairing occurs that can be represented by the following equations:



A pairing scheme such as this is not unreasonable because the alkali metals exist as diatomic molecules in the vapor state. In concentrated solutions, it is believed that metal “ions” and electrons form some sort of continuum similar to a bulk metal with regard to its properties as an electrical conductor.

As might be expected for solutions containing free electrons, solutions of metals in ammonia are not stable over long time periods, and hydrogen is slowly liberated.



These solutions behave as strong reducing agents, in accord with their containing “free” electrons. For example, they will react with oxygen to produce superoxide and peroxide ions as shown by the equations



It is also possible to prepare compounds that contain metals in unusual oxidation states. For example, the reaction of  $\text{K}_2[\text{Ni}(\text{CN})_4]$  with potassium in liquid ammonia produces  $\text{K}_4[\text{Ni}(\text{CN})_4]$  as the nickel is reduced from +2 to 0.

To a considerable extent, solutions of alkali metals in solvents such as methyl amine or ethylenediamine exhibit similar properties to the ammonia solutions. Some of the alkali metals also dissolve in ethers and tetrahydrofuran, but that chemistry will not be described here.

### 6.2.5 Liquid HF

Although liquid HF attacks many materials, including glass, it has been extensively studied as a nonaqueous solvent. Because of strong hydrogen bonding (see Chapter 4), HF is a liquid over the range of temperatures from  $-83.1$  to  $19.5$  °C. In this regard, it is somewhat more convenient to work with than  $\text{NH}_3$ . However, containers must usually be made of an inert material such as Teflon (polytetrafluoroethylene). Some of the relevant physical properties of HF are summarized in Table 6.7.

The equivalent conductance of HF ( $1.4 \times 10^{-5} \text{ ohm}^{-1}$ ) is higher than that of water ( $6.0 \times 10^{-8} \text{ ohm}^{-1}$  at  $25$  °C) indicating a somewhat larger degree of autoionization that can be illustrated by the reaction



Because of its high dielectric constant and dipole moment, HF is a good solvent for many inorganic salts with many compounds being more soluble in liquid HF than in water. Organic compounds also dissolve in HF and the acidic nature of the solvent enables it to catalyze many reactions. The cation characteristic of the solvent,  $\text{H}_2\text{F}^+$ , is such a strong acid that only a few compounds react with HF to produce it. To some extent, the acidic character is enhanced when the acid is capable of forming stable fluoro complexes, as in the following cases.



In these equations, the acid species present are sometimes represented as  $\text{HBF}_4$  and  $\text{HSbF}_6$ , respectively.

The amphoteric behavior of  $\text{Zn}^{2+}$  in water and ammonia has been shown in Eqs (6.70)–(6.74). Similar behavior can be shown for  $\text{Al}^{3+}$  in liquid HF because  $\text{AlF}_3$  is insoluble in HF. The reactions of  $\text{Al}^{3+}$  in liquid HF can be shown as follows.



Note that in Eq. (6.105) the addition of NaF represents the addition of a base in HF because  $\text{F}^-$  (or  $\text{HF}_2^-$  if we write it as being solvated) is the anion characteristic of the solvent.

Some organic reactions, particularly fluorination reactions, can be carried out in HF solutions. Also, because nitration is an acid-catalyzed reaction, such reactions can be accomplished in HF. There are numerous other aspects to the chemistry of liquid HF that will not be described in this survey.

**TABLE 6.7** Physical Properties of HF

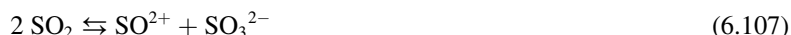
Melting point	$-83.1$ °C
Boiling point	$19.5$ °C
Density at $19.5$ °C	$0.991 \text{ g cm}^{-3}$
Heat of fusion	$4.58 \text{ kJ mol}^{-1}$
Heat of vaporization	$30.3 \text{ kJ mol}^{-1}$
Equivalent conductance	$1.4 \times 10^{-5} \text{ ohm}^{-1}$
Dielectric constant at $0$ °C	83.6
Dipole moment	1.9 D

## 6.2.6 Liquid Sulfur Dioxide

Although no attempt will be made to describe the chemistry of all of the nonaqueous solvents listed in Table 6.4, the survey to this point has included ammonia as a basic solvent and liquid HF as an acidic solvent. Another solvent that has been extensively utilized in both inorganic and organic chemistry is sulfur dioxide. Accordingly, we will give a brief survey of the chemistry of liquid sulfur dioxide for which the physical properties are presented in Table 6.8.

Although the dipole moment of  $\text{SO}_2$  is fairly high, the dielectric constant is much lower than that of water and slightly lower than that of ammonia. Consequently, liquid  $\text{SO}_2$  is a poorer solvent for most ionic salts than is water, HF, or  $\text{NH}_3$ . Covalent compounds are generally more soluble in liquid  $\text{SO}_2$  than in water, and liquid sulfur dioxide is a good solvent for many such materials. This is, of course, in accord with the prediction made on the basis of the HSIP. The solvents containing smaller, harder molecules such as water and HF are better solvents for ionic compounds. On the other hand, liquid  $\text{SO}_2$  is a very good solvent for a wide range of organic compounds. However, aromatic compounds are much more soluble than are aliphatic compounds, and this difference has been utilized as the basis for a solvent extraction process for separating aliphatic and aromatic hydrocarbons.

The low conductivity of liquid  $\text{SO}_2$  has been interpreted as arising from ions produced by autoionization.



However, generating doubly charged ions would be highly unfavorable compared to autoionization of water, ammonia, or HF in which proton transfer occurs. As a result, if autoionization does occur, the extent must be very slight. In fact, there is no experimental basis for assuming that any autoionization occurs. Radioactive sulfur in  $\text{SOCl}_2$  is not exchanged with the sulfur in liquid  $\text{SO}_2$ . If the two solvents undergo ionization to produce  $\text{SO}^{2+}$ , it would be expected that exchange of sulfur would occur.

Although it is unlikely that  $\text{SO}_2$  undergoes ionization as shown in Eq. (6.107), some reactions take place as though these ions were present. For example, because according to the solvent concept  $\text{SO}^{2+}$  would be the acidic species in liquid  $\text{SO}_2$ , a compound such as  $\text{SOCl}_2$  would be expected to react as an acid. Similarly,  $\text{K}_2\text{SO}_3$  would be a base because it would provide the basic  $\text{SO}_3^{2-}$  ions. Therefore, the reaction between the acid and base in liquid  $\text{SO}_2$  would be



However, this does not mean that neutralization of  $\text{SOCl}_2$  with  $\text{K}_2\text{SO}_3$  can be represented by the ionic equation



It should be noted that the products of Eq. (6.108) are exactly the same as if ionization had occurred. The solvent concept has some utility even in cases where it does not exactly represent the way in which the solvent behaves with regard to autoionization.

Presuming the solvent concept to apply in the case of liquid sulfur dioxide,  $\text{SOCl}_2$  would be an acid that produces  $\text{SO}^{2+}$ . It may be that  $\text{SOCl}_2$  undergoes some slight autoionization that can be represented as

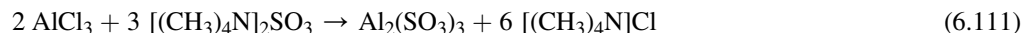


**TABLE 6.8** Physical Properties of  $\text{SO}_2$

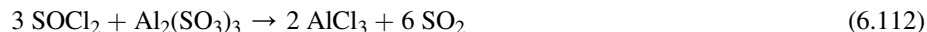
Melting point	$-75.5^\circ\text{C}$
Boiling point	$-10.0^\circ\text{C}$
Density	$1.46 \text{ g cm}^{-3}$
Heat of fusion	$8.24 \text{ kJ mol}^{-1}$
Heat of vaporization	$24.9 \text{ kJ mol}^{-1}$
Dipole moment	1.61 D
Dielectric constant	15.6
Specific conductance	$3 \times 10^{-8} \text{ ohm}^{-1}$

although  $\text{Cl}^-$  would presumably be present as a solvated species,  $\text{SOCl}_3^-$ . This behavior is suggested by the fact that radioactive chlorine from other soluble chlorides exchanges with the chlorine atoms from  $\text{SOCl}_2$ .

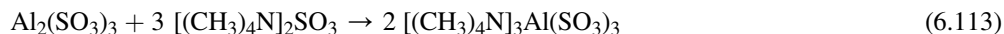
Amphoteric behavior in liquid  $\text{SO}_2$  is similar to that in other solvents. In liquid  $\text{SO}_2$ , aluminum sulfite is relatively insoluble, which leads to the reaction



Subsequently, aluminum sulfite reacts with an acid such as  $\text{SOCl}_2$ ,



and with a base such as  $[(\text{CH}_3)_4\text{N}]_2\text{SO}_3$ ,



These equations are analogous to those shown earlier for amphoteric behavior of  $\text{Zn}^{2+}$  and  $\text{Al}^{3+}$  in water, liquid ammonia, and liquid HF.

In describing some of the chemistry of nonaqueous solvents, we have focused on reactions in which the solvent itself is a reactant. Many other reactions occur in which the solvent is only the reaction medium. These reactions are often similar regardless of the solvent chosen. Consider, for example, the following reactions:



Numerous examples of similar metathesis reactions could be cited, but in each case, one product is insoluble in the solvent. Reactions such of these are not restricted to a particular solvent.

We have chosen to briefly describe the behavior of three representative nonaqueous solvents. As shown in Table 6.4, numerous other compounds have been utilized as nonaqueous solvents, and the chemistry of some of them will be described in the chapters dealing with the chemistry of the specific elements (e.g.,  $\text{BrF}_3$  in Chapter 17 that deals with halogen chemistry). Nonaqueous solvents constitute a vast area of descriptive inorganic chemistry, which has briefly been introduced because of the space limitations of this book, not because of the importance of the subject.

### 6.3 SUPERACIDS

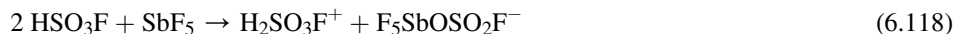
As a consequence of the leveling effect, the strongest acid that can exist in aqueous solutions is  $\text{H}_3\text{O}^+$ . However, there exists a class of acidic materials that exceed the acidity of  $\text{H}_3\text{O}^+$ , and they are referred to as *superacids*. These species are generated by altering the structure of an acid such as  $\text{H}_2\text{SO}_4$  to increase the inductive effect on an O–H bond. Such a material may exhibit an acid strength that is as much as  $10^{10}$ – $10^{15}$  times that of  $\text{H}_2\text{SO}_4$  alone. One such compound is  $\text{HSO}_3\text{F}$  (m. p.  $-89^\circ\text{C}$ , b. p.  $163^\circ\text{C}$ ) in which the inductive effect of the fluorine atom is greater than that which a second OH group would produce. As a result, it is a stronger acid than  $\text{H}_2\text{SO}_4$ .

A mixture of HF and  $\text{SbF}_5$  also functions as a superacid as a result of the reaction



This mixture can be utilized in solvents such as  $\text{SO}_2$ ,  $\text{SO}_2\text{FCl}$ , and  $\text{SO}_2\text{F}_2$ , but it is somewhat weaker than  $\text{H}_2\text{SO}_3\text{F}^+$ .

When  $\text{SbF}_5$  is added to  $\text{HSO}_3\text{F}$ , the very strong Lewis acidity of  $\text{SbF}_5$  results in an acid that is even stronger than  $\text{HSO}_3\text{F}$  as a result of the reaction



The oxidizing agent peroxydisulfuryl difluoride,  $\text{FO}_2\text{SOOSO}_2\text{F}$ , can be used to carry out reactions such as the production of halogen cations. One such process can be shown as



In such a strong oxidizing medium it is also possible to produce polyatomic cations of sulfur.



Even hydrocarbon molecules can react to produce cations as illustrated by the following.



The use of superacids in nonaqueous solvents makes possible many reactions that could not be carried out in any other way. This is one of the aspects of the chemistry of nonaqueous solvents that makes this area so important.

## REFERENCES FOR FURTHER READING

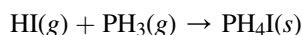
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## PROBLEMS

- Arrange the following in the order of increasing basicity and explain your reasoning:  $\text{CO}_2$ ,  $\text{Br}_2\text{O}_5$ ,  $\text{CaO}$ ,  $\text{Cl}_2\text{O}_3$ .
- Give the conjugate base of each of the following:
  - $\text{OH}^-$
  - $\text{HNO}_2$
  - $\text{HSe}^-$
  - $\text{PH}_3$
  - $\text{H}_2\text{AsO}_4^-$
- Give the conjugate acid of each of the following:
  - $\text{OH}^-$
  - $\text{NH}_3$
  - $\text{CN}^-$
  - $\text{HPO}_4^{2-}$
  - $\text{HS}^-$
- Which of the following is the strongest Brønsted base? Explain your answer.  $\text{S}^{2-}$ ,  $\text{O}^{2-}$ ,  $\text{F}^-$ , or  $\text{I}^-$
- Consider the acids  $\text{HNO}_3$ ,  $\text{HClO}_3$ ,  $\text{H}_3\text{BO}_3$ ,  $\text{HBrO}_4$ .
  - Which would be the strongest acid?
  - Which would be the weakest acid?
  - Which two would be of comparable strength?
- Each of the following reactions goes more than 50% toward completion.
  - $\text{HSO}_4^- + \text{C}_2\text{H}_3\text{O}_2^- \rightarrow \text{HC}_2\text{H}_3\text{O}_2 + \text{SO}_4^{2-}$
  - $\text{HC}_2\text{H}_3\text{O}_2 + \text{HS}^- \rightarrow \text{C}_2\text{H}_3\text{O}_2^- + \text{H}_2\text{S}$
  - $\text{HCO}_3^- + \text{OH}^- \rightarrow \text{CO}_3^{2-} + \text{H}_2\text{O}$

For each of the equations above, list of all the Brønsted acids present and arrange them in the order of strongest to weakest.
- For each of the following pairs, select the stronger Brønsted acid and explain your answer.
  - $\text{H}_3\text{BO}_3$  or  $\text{H}_2\text{CO}_3$
  - $\text{HSO}_4^-$  or  $\text{HCO}_3^-$
  - $\text{H}_2\text{Se}$  or  $\text{HI}$
  - $\text{HOCl}$  or  $\text{HCl}$
  - $\text{HNO}_2$  or  $\text{HPO}_3$

8. Write a complete, balanced equation to show what would happen when each of the following substances is added to water.
- Acetic acid
  - Trimethylamine,  $(\text{CH}_3)_3\text{N}$
  - $\text{NH}_4\text{NO}_3$
  - $\text{NaOCl}$
  - $\text{KC}_2\text{H}_3\text{O}_2$
9. Account for the fact that phosphorous acid,  $\text{H}_3\text{PO}_3$ , is a weak, diprotic acid whereas hypophosphorous acid,  $\text{H}_3\text{PO}_2$ , is a weak, monoprotic acid.
10. Consider the acids  $\text{H}_3\text{CCOOH}$ ,  $\text{F}_3\text{CCOOH}$ , and  $\text{Cl}_3\text{CCOOH}$ . Arrange these in the order of decreasing strength and explain your answer.
11. Each of the following reactions proceeds more than 50% toward completion. On the basis of these reactions, arrange all of the Brønsted acids in the order of decreasing strength.
- $\text{H}_3\text{PO}_4 + \text{N}_3^- \rightarrow \text{HN}_3 + \text{H}_2\text{PO}_4^-$
  - $\text{HN}_3 + \text{OH}^- \rightarrow \text{H}_2\text{O} + \text{N}_3^-$
  - $\text{H}_3\text{O}^+ + \text{H}_2\text{PO}_4^- \rightarrow \text{H}_3\text{PO}_4 + \text{H}_2\text{O}$
  - $\text{H}_2\text{O} + \text{PH}_2^- \rightarrow \text{PH}_3 + \text{OH}^-$
12. Write complete equations to show the reaction of each of the following with water.
- $\text{NO}_2^-$
  - $\text{CN}^-$
  - $\text{N}_2\text{H}_5^+$
  - $\text{HCO}_3^-$
13. Write a complete, balanced equation to show what would happen when each of the following substances is added to water.
- $\text{Na}_2\text{CO}_3$
  - $(\text{CH}_3)_3\text{NHCl}$
  - $\text{CaH}_2$
  - $\text{NaHSO}_4$
  - $\text{NaH}_2\text{PO}_4$
14. Explain why  $\text{PH}_3$  is a much weaker base in water than is  $\text{NH}_3$ .
15. Tell whether aqueous solutions of the following will be acidic, basic, or neutral. For those that are not neutral, write an equation to show why they are not.
- $\text{NH}_4\text{NO}_3$
  - $\text{NaNO}_3$
  - $\text{NaHCO}_3$
  - $\text{K}_3\text{AsO}_4$
  - $\text{KCN}$
16. Explain why water is a better medium than liquid ammonia for a reaction that requires an acidic medium although the opposite is true for a reaction that requires a basic medium.
17. Explain why the reaction



is an acid base reaction according to (a) the Brønsted theory and (b) the Lewis theory.

18. On the basis of the HSAB principle, which of these reactions will likely take place?
- $\text{AgF}_2^- + 2 \text{Br}^- \rightarrow \text{AgBr}_2^- + 2 \text{F}^-$
  - $\text{Ni}(\text{CO})_4 + 4 \text{NH}_3 \rightarrow \text{Ni}(\text{NH}_3)_4 + 4 \text{CO}$
  - $(\text{C}_2\text{H}_5)_2\text{O}:\text{BCl}_3 + (\text{C}_2\text{H}_5)_2\text{S}:\text{S} \rightarrow (\text{C}_2\text{H}_5)_2\text{S}:\text{BCl}_3 + (\text{C}_2\text{H}_5)_2\text{O}:\text{S}$
  - $\text{PH}_4\text{I} + \text{NH}_3 \rightarrow \text{NH}_4\text{I} + \text{PH}_3$

19. On the basis of the hard–soft interaction principle, predict which species would be more stable and explain your answer.
- (a)  $\text{Fe}(\text{CO})_5$  or  $\text{Fe}(\text{NH}_3)_5$
  - (b)  $\text{Cr}(\text{CO})_6^{3+}$  or  $\text{Cr}(\text{NH}_3)_6^{3+}$
  - (c)  $\text{Pt}(\text{O}(\text{C}_2\text{H}_5)_2)_4^{2+}$  or  $\text{Pt}(\text{S}(\text{C}_2\text{H}_5)_2)_4^{2+}$
20. Explain why the bonding of  $\text{SO}_3^{2-}$  is different to  $\text{Cr}^{3+}$  than it is to  $\text{Pt}^{2+}$ .
21. Explain why  $\text{NF}_3$  is a weaker Lewis base than  $\text{NH}_3$ .
22. Explain the difference in acidity of  $\text{HCN}$  in water and liquid ammonia.
23. Explain why there are no hydride complexes of the type  $\text{Co}(\text{NH}_3)_5\text{H}^{2+}$ , but there are some such as  $\text{Mo}(\text{CO})_5\text{H}$  that do formally contain  $\text{H}^-$  as a ligand.
24. Assume that pure acetic anhydride,  $(\text{CH}_3\text{CO})_2\text{O}$ , undergoes some autoionization.
- (a) What species would be produced? Write an equation to represent the process.
  - (b) Write the equations that explain why an aluminum ion is amphoteric in liquid acetic anhydride.
25. Write an equation for the process indicated.
- (a) The neutralization of  $\text{N}_2\text{H}_4$  with sulfuric acid
  - (b) The reaction of  $\text{CO}_2$  with a solution of  $\text{Ca}(\text{OH})_2$
  - (c) The reaction of  $\text{CaO}$  with  $\text{SnO}_2$  at high temperature
  - (d) The reaction of  $\text{N}_2\text{O}_5$  with water
  - (e) The reaction of  $\text{Cl}_2\text{O}$  with water
26. The reaction of  $\text{AlCl}_3$  with  $\text{OCCl}_2$  results in the formation of some  $\text{AlCl}_4^-$ .
- (a) Write an equation showing how  $\text{AlCl}_4^-$  forms assuming that the solvent concept applies to this system.
  - (b) Assume that the coordination model applies to this system and write an equation(s) to show how  $\text{AlCl}_4^-$  is formed.
  - (c) Explain clearly the difference between the solvent concept and the coordination model as related to the reaction of  $\text{AlCl}_3$  with  $\text{OCCl}_2$ .
27. Which is a more stable solid,  $\text{PH}_4\text{F}$  or  $\text{NH}_4\text{F}$ ? Why?
28. Explain how  $\text{KNO}_3$  can be used as a nitrating agent in liquid  $\text{HF}$ . What species are present? How do they react?
29. When sodium is dissolved in liquid ammonia to prepare a dilute solution, the solution is less dense than liquid ammonia alone and the solution is paramagnetic. Explain these observations.



# Hydrogen

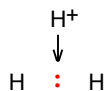
Although hydrogen is the simplest atom, the chemistry of hydrogen is extensive. Part of the reason for this is because the hydrogen atom resembles the members of two groups of elements. Because it has the electron configuration of  $1s^1$  it resembles to some extent the alkali metals in Group IA, which have configurations of  $ns^1$ . However, the ionization energy of hydrogen is about  $1314 \text{ kJ mol}^{-1}$  and the strength of the H–H bond in the diatomic molecule is about  $435 \text{ kJ mol}^{-1}$ . Consequently, the formation of a simple  $\text{H}^+$  species requires considerable energy. In fact, so much energy is required for the formation of  $\text{H}^+$  that compounds in which the single proton is present as a cation are not likely to form. In contrast, the alkali metals have ionization energies varying from about  $377 \text{ kJ mol}^{-1}$  for Cs to about  $519 \text{ kJ mol}^{-1}$  for Li so that many compounds of these elements are essentially ionic and contain the singly charged ions. There are, however, some solid compounds that contain ions such as  $\text{H}_5\text{O}_2^+$  or  $\text{H}_9\text{O}_4^+$ . These ions are solvated protons that contain two and four water molecules of hydration, respectively. The reason for the existence of ions of this type is the high heat of hydration of  $\text{H}^+$  ( $-1100 \text{ kJ mol}^{-1}$ ), which results from its small size and the resulting high charge to size ratio. The heat of hydration of  $\text{H}_3\text{O}^+$  is estimated to be about  $-390 \text{ kJ mol}^{-1}$ .

Hydrogen was discovered by Cavendish in 1781 who prepared a gas that produces water during its combustion. Shortly thereafter, the name hydrogen was given to the gas. Most simple hydrogen compounds are covalent as a result of the hydrogen atom sharing an electron pair. Because the  $1s$  level is singly occupied in the hydrogen atom, it also resembles in many ways the halogens, which also require a single electron to complete the valence shell. Accordingly, hydrogen forms a substantial number of compounds in which it gains an electron to form a hydride ion. Thus, although hydrogen is the simplest atom, the chemistry of hydrogen is indeed varied and encompasses many types of compounds and reactions.

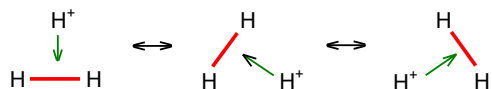
## 7.1 ELEMENTAL AND POSITIVE HYDROGEN

As a result of the nucleus of the hydrogen atom being a proton that has a spin quantum number of  $1/2$ , a hydrogen molecule may have the spins both being aligned or opposed. The result is that there are two forms of elemental hydrogen. These are known as ortho  $\text{H}_2$  if the spins are aligned or para hydrogen if the spins are opposed. At room temperature, the mixture is composed of 75% ortho and 25% para hydrogen.

The absence of compounds that contain the simple  $\text{H}^+$  ion has been discussed above. However, the solvated species are not the only species known that contain  $\text{H}^+$  in some form. A series of cations have been identified in mass spectrometry that result from the attachment of a hydrogen molecule to  $\text{H}^+$ . The simplest of these is the  $\text{H}_3^+$  ion that has a trigonal planar structure. The electrostatic interaction of a proton with  $\text{H}_2$  occurs at the shared pair of electrons that can be shown as follows.

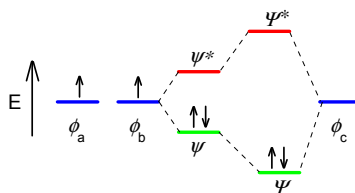


After  $\text{H}^+$  attaches, the electrons (and there are only two) are shared equally in an equilateral triangular arrangement. Of course there are multiple resonance structures possible, which can be shown as follows.



The  $\text{H}_3^+$  ion represents the simplest example of a two-electron three-center bond in which a molecular orbital containing two electrons encompasses all three of the atoms. Instead of the approach described above, a more satisfactory description of the bonding is provided by constructing a molecular orbital from a combination of three hydrogen wave functions,

$$\psi_{\text{MO}} = \frac{1}{\sqrt{3}}(\phi_a + \phi_b + \phi_c) \quad (7.1)$$



**FIGURE 7.1** A molecular orbital diagram for  $\text{H}_3^+$  in which  $\phi_a$ ,  $\phi_b$ , and  $\phi_c$  represent  $1s$  atomic wave functions,  $\psi$  and  $\psi^*$  are wave functions for bonding and antibonding states in  $\text{H}_2$ , and  $\Psi$  and  $\Psi^*$  are wave functions for the  $\text{H}_3^+$  ion.

where  $\psi_{\text{MO}}$  is the wave function for the molecular orbital and  $\phi_a$ ,  $\phi_b$ , and  $\phi_c$  represent the  $1s$  atomic wave functions for hydrogen atoms a, b, and c, respectively. The resulting molecular orbital diagram can be represented as shown in Figure 7.1.

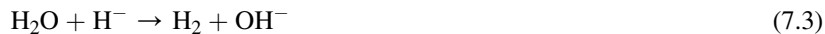
In this case, it is assumed that the molecular orbitals in a hydrogen molecule can be represented in the usual way (see Section 3.3.1) and that the bonding molecular orbital is then used with the third atomic orbital to give the bonding molecular orbital for the entire species. Because  $\text{H}_3^+$  contains only two electrons, the bonding molecular orbital is occupied by two electrons, but it encompasses all three atomic centers. This type of three-center bonding will also be discussed in later chapters.

The series of species containing a proton attached to hydrogen molecules can be represented by the formula  $\text{H}_n^+$  with  $\text{H}_3^+$  being the most stable, and others having  $n$  being an odd number are significantly more stable than those in which  $n$  is an even number. These species have structures that are shown in Figure 7.2. The ions arise from the addition of  $\text{H}_2$  molecules at the corners of the triangle formed in  $\text{H}_3^+$ .

Elemental hydrogen can be prepared conveniently by either the reduction of positive hydrogen, as typified by the reactions of metals with mineral acids, or the oxidation of the hydride ion. The first process may be represented as



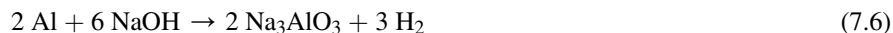
where M is a metal above hydrogen in the electromotive series and  $\text{A} = \text{Cl}^-$ ,  $\text{Br}^-$ ,  $1/2 \text{SO}_4^{2-}$ , etc. The oxidation of hydride ion by positive hydrogen is the simplest reaction yielding hydrogen from hydrides. The following are typical reactions of this type, but they also represent proton transfer (acid–base) processes.



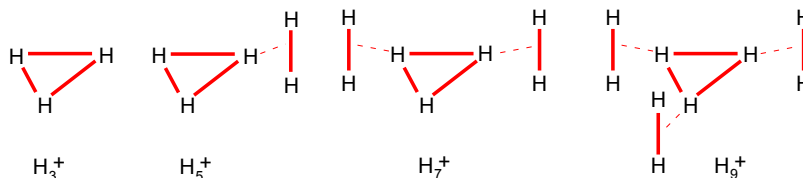
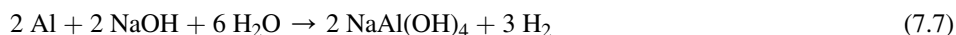
Elemental hydrogen may also be prepared by the electrolysis of water.



Because the reduction of hydrogen at the cathode can be essentially an isolated process, this method is useful for the production of high purity  $\text{H}_2$ . Hydrogen is also liberated by the action of  $\text{NaOH}$  on amphoteric metals such as aluminum and zinc.



In aqueous solution, this reaction can be represented more accurately by the equation



**FIGURE 7.2** Structures of some  $\text{H}_n^+$  ions.

To be commercially feasible, major industrial preparations of hydrogen must use inexpensive materials and processes. A reducing agent that is as expensive as a metal or electricity simply will not give an economically feasible process for preparing hydrogen. At very high temperatures, carbon in the form of coke (obtained by heating coal) reacts with water to produce  $\text{H}_2$ .



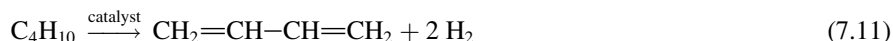
The mixture of CO and  $\text{H}_2$  is known as *water gas*. It is a good reducing agent (both CO and  $\text{H}_2$  are easily oxidized), and it has been used as a fuel because both gases will readily burn. Large quantities of hydrogen are also produced by treatment of petroleum products by either catalytic reforming or dehydrogenation. In catalytic reforming, hydrocarbons have hydrogen removed, and structural changes are produced. For example, hexane is converted to cyclohexane.



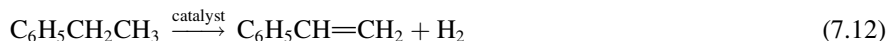
The cyclohexane then can be converted to benzene.



The catalysts used in these types of processes are usually platinum in some form. Dehydrogenation involves removal of hydrogen, and two important processes are conversion of butane to butadiene,



and the conversion of ethyl benzene to styrene,



Both butadiene and styrene are used in very large quantities in the preparation of polymers so these dehydrogenation processes result in the production of large quantities of hydrogen.

One of the most important processes for the production of hydrogen is the steam reformer process. In that process, a hydrocarbon and steam are passed over a nickel catalyst at  $900^\circ\text{C}$ . The reactions that take place are shown below when the hydrocarbon is methane.



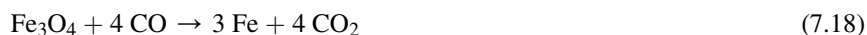
The mixture of gases that results contains  $\text{H}_2$ , CO, and  $\text{CO}_2$  along with some steam. This mixture is then passed into a shift converter and at  $450^\circ\text{C}$ , the CO is converted into  $\text{CO}_2$ .



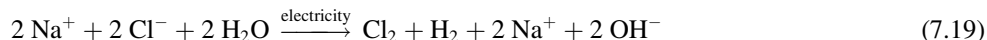
In years past, another process for making hydrogen was important. The basis for that process was the following reaction.



When all the iron becomes oxidized, it can be regenerated to Fe by the following reactions carried out at red heat.



Finally, a preparation of hydrogen that is of considerable importance is the electrolysis of aqueous NaCl solutions.



This process is of enormous importance because it represents the commercial preparation of chlorine and sodium hydroxide. Because it is carried out on a very large scale, this process also produces large quantities of hydrogen.

Hydrogen is used as the fuel in the oxy–hydrogen torch to produce an extremely hot flame. It is also used in bubble chambers for tracking elementary particles and as a rocket fuel. Hydrogen has several attributes that make it an attractive

energy source. First, a great deal of energy is released when hydrogen burns. Second, the combustion product is water so there are no environmental issues. Third, the major source of hydrogen is water, which is abundantly available. However, the difficulty is how to separate the  $\text{H}_2\text{O}$  molecules to obtain large quantities of hydrogen economically. There are also difficulties associated with fabricating engines that run for extended periods because of the reaction of hydrogen with metals (see Section 7.3.2) and with storage of hydrogen in the liquid state. Thus, an economy based on hydrogen as a major fuel may not be forthcoming soon.

## 7.2 OCCURRENCE AND PROPERTIES

Elemental hydrogen does not occur in the earth's atmosphere to any significant extent owing to the low molecular mass of the molecules. However, it does occur to a large extent in other parts of the universe. Combined hydrogen is present on the earth's surface in a wide variety of compounds, especially water. The principal use of elemental hydrogen is in several important hydrogenation reactions, notably in the production of ammonia, methyl alcohol and a large number of organic materials and foodstuffs.

Ordinary hydrogen consists of a mixture of three isotopes. The isotope with mass number 1, H, is about 6400 times as abundant as deuterium, D, the isotope with mass number 2. Tritium, T, the isotope with mass number 3, is many orders of magnitude less abundant than deuterium. Although six different diatomic molecules are possible from these three isotopes,  $\text{H}_2$ ,  $\text{D}_2$ , HD, and  $\text{T}_2$  have been more thoroughly studied. Of these, the two most common forms are  $\text{H}_2$  and  $\text{D}_2$ . Some of the properties of these two forms of hydrogen are listed in Table 7.1.

Although  $\text{D}_2$  undergoes almost all the same reactions as  $\text{H}_2$ , the rates of these reactions are generally lower. In some cases, there is a large difference between the rate of the reaction when deuterium is involved and the rate when hydrogen is involved. The rate is influenced by the kinetic isotope effect, which refers to the fact that the atomic mass of deuterium being twice that of hydrogen causes the rates of most reactions involving deuterium to be lower. It is only in the case of hydrogen and deuterium that such a large relative mass difference occurs. The effect arises from the fact that whereas X–H and X–D bonds are about the same strength, the vibrational frequencies are much different with that for X–H being much higher. Therefore, the X–H bond reacts more rapidly.

By far the most frequently encountered deuterium compound is  $\text{D}_2\text{O}$  or “heavy” water. This material is generally obtained by enrichment from natural water by electrolysis. The normal water,  $\text{H}_2\text{O}$ , is preferentially electrolyzed leaving water enriched in  $\text{D}_2\text{O}$ . Eventually, almost pure  $\text{D}_2\text{O}$  can be obtained. Some of the properties of  $\text{D}_2\text{O}$  and  $\text{H}_2\text{O}$  are shown in Table 7.2.

**TABLE 7.1** Physical Properties of  $\text{H}_2$  and  $\text{D}_2$

Property	$\text{H}_2$	$\text{D}_2$
Boiling point, K	20.28	23.59
Triple point, K	13.92	18.71
Heat of fusion, $\text{J mol}^{-1}$	117	219

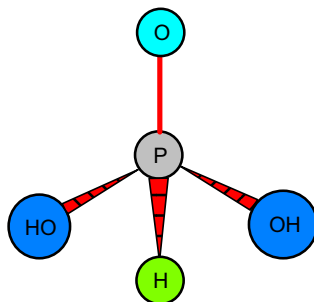
**TABLE 7.2** Properties of  $\text{H}_2\text{O}$  and  $\text{D}_2\text{O}$

Property	$\text{H}_2\text{O}$	$\text{D}_2\text{O}$
Melting point, $^{\circ}\text{C}$	0.00	3.8
Boiling point, $^{\circ}\text{C}$	100.00	101.42
Temperature of maximum density, $^{\circ}\text{C}$	3.96	11.6
Heat of vaporization, $\text{kJ mol}^{-1}$	40.66	41.67
Heat of fusion, $\text{kJ mol}^{-1}$	6.004	6.368
Dielectric constant at $20^{\circ}\text{C}$	82	80.5
Solubility of KCl at $0^{\circ}\text{C}$ , moles per 100 mol of solvent	6.81	5.69

One of the important features of  $D_2O$  is its ability to undergo isotope exchange with many compounds. For example, when  $NH_3$  is placed in  $D_2O$ , all of the H atoms exchange rapidly with deuterium. Presumably, this is because the exchange with  $NH_3$  can proceed by interaction of  $D_2O$  with the unshared pair of electrons on the nitrogen atom in an associative process that can be illustrated as



The exchange with  $NH_4^+$  is much slower because attachment of deuterium depends on a dissociative mechanism of  $NH_4^+$ . The exchange of D with H in  $[Co(NH_3)_6]^{3+}$  in  $D_2O$  is very slow for the same reason. Likewise, the exchange of hydrogen by formic acid,  $HCOOH$ , in  $D_2O$  proceeds rapidly for the hydrogen attached to oxygen, but very slowly for the hydrogen attached to carbon. For phosphorous acid,  $H_3PO_3$ , two hydrogen atoms exchange rapidly, but the third does not. The reason for this is that in this molecule, one of the hydrogen atoms is bonded directly to the phosphorus atom.



In order to undergo facile exchange between D and H, the bond must have some polarity. Because the electronegativities of P and H are almost exactly the same, the hydrogen atom attached to the phosphorus atom does not exchange readily. As we have seen above, even if the bond is polar (as in  $NH_3$ ), there must be an unshared electron pair to provide a site of attack or at least there must be some low energy pathway for the reaction.

The major use of  $D_2O$  is as a moderator for neutrons in nuclear reactors. Because the deuterium atom is small, neutrons colliding with it cause the deuterium atom to recoil so that more energy is absorbed than would be if the atom did not recoil. Thus, “heavy” water is used for this purpose. Most of the other compounds of interest that contain positive hydrogen will be discussed in chapters on the chemistry of the other elements. We will now discuss the chemistry of negative hydrogen.

## 7.3 HYDRIDES

In most compounds, hydrogen exhibits the oxidation state of +1. Because of its relatively high electronegativity, there are compounds in which hydrogen assumes a negative oxidation state either by gaining an electron in the  $1s$  state or by sharing electrons with an element of lower electronegativity. Formally, this is true in any binary compound in which the other element has a lower electronegativity than hydrogen. Hydrides are conveniently grouped into three classes, depending upon the mode of binding to the other element. However, as in the case of other chemical bonds, the bond types are not completely separable because there is a gradual transition from ionic to covalent depending on the relative electronegativities of the atoms. As a result, the three classifications of hydrides described below are somewhat artificial.

### 7.3.1 Ionic Hydrides

It is logical to expect that elements of low electronegativity might produce hydride ions by giving up electrons to hydrogen. Hydrogen has an electron affinity of  $-74.5 \text{ kJ mol}^{-1}$ , but this is not sufficient to overcome the ionization energy of the metal if only the pair of ions were formed. The process becomes energetically favorable when a solid crystal is formed, and solid hydrides are known for the Groups IA and IIA elements of the periodic table. Generally, these compounds can be formed by direct union of the elements.



The reaction is carried out at about  $400^\circ\text{C}$  for sodium, potassium, and rubidium, and at about  $700^\circ\text{C}$  with lithium. Metal compounds such as the nitrides may also be used.



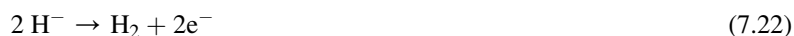
**TABLE 7.3** Selected Properties of the Group I Hydrides

Compound	$\Delta H_f^\circ$ (kJ mol <sup>-1</sup> )	$U$ (kJ mol <sup>-1</sup> )	H <sup>-</sup> Radius (pm)	Apparent Charge on H (e units)	Density Hydride (g cm <sup>-3</sup> )	Density Metal (g cm <sup>-3</sup> )
LiH	-89.1	916	136	-0.49	0.77	0.534
NaH	-59.6	808	146	-0.50	1.36	0.972
KH	-57.7	720	152	-0.60	1.43	0.859
RbH	-47.7	678	153	-0.63	2.59	1.525
CsH	-42.6	644	154	-0.65	3.41	1.903

The ionic hydrides are white solids with high melting points, and all of the alkali metal hydrides have the sodium chloride crystal structure. Because they resemble the salts of the alkali and alkaline earth metals, the ionic hydrides are often referred to as saline or saltlike hydrides. The properties of the alkali metal hydrides are shown in Table 7.3 and those of the alkaline earth hydrides are shown in Table 7.4.

The apparent radius of the H<sup>-</sup> is determined from the M–H distance in the crystal by subtracting the known radius of M<sup>+</sup>. It is obvious from the data shown in Table 7.3 that the radius assigned to H<sup>-</sup> in Li–H is smaller than in other hydrides of Group IA. The Li–H bond is considered to have a substantial amount of covalent character. Probably the greater covalency in this compound compared to the other alkali metal hydrides is a result of the higher ionization potential of Li and the fact that the 1s orbital of hydrogen and the 2s orbital of lithium are of similar size. The vast difference in size between 1s of H and the 3s or 4s of Na or K decreases the effectiveness of overlap. Also, the fact that Na and K have lower ionization potentials than Li tends to make NaH and KH more ionic than LiH. The H<sup>-</sup> ion is a large, soft electronic species that is easily polarizable so electron density would be drawn toward a smaller ion such as Li<sup>+</sup> to a greater extent than with other ions of Group IA metals.

Molten hydrides of Groups IA and IIA are good electrical conductors, and hydrogen is liberated at the anode as a result of the oxidation of H<sup>-</sup>.



The alkali metal or alkaline earth metal is reduced at the cathode.

The dominant feature of the chemistry of the ionic hydrides is the strongly basic character of the H<sup>-</sup> ion. All the ionic hydrides react readily with protonic solvents to produce hydrogen gas and a base that is weaker than H<sup>-</sup>.



Because the hydride ion has a pair of electrons, it can function as a Lewis base as well. For example, many coordination compounds are known in which the hydride ion is present as a ligand. Some of the simplest of these are the tetrahydroaluminate(III), AlH<sub>4</sub><sup>-</sup>, and the tetrahydridoborate(III) ions. Although these species are largely covalent, they can be considered as arising from the coordination of four H<sup>-</sup> ions to Al<sup>3+</sup> and B<sup>3+</sup> ions. The salts of these complex ions, LiAlH<sub>4</sub>

**TABLE 7.4** Some Properties of Alkaline Earth Hydrides

Compound	$\Delta H_f^\circ$ (kJ mol <sup>-1</sup> )	Apparent Charge on H (e units)	Density, Hydride (g cm <sup>-3</sup> )	Density, Metal (g cm <sup>-3</sup> )
CaH <sub>2</sub>	-195	-0.27	1.90	1.55
SrH <sub>2</sub>	-177	-0.31	3.27	2.60
BaH <sub>2</sub>	-172	-0.36	4.15	3.59

and  $\text{NaBH}_4$ , are widely used as hydrogenating (reducing) agents in organic synthesis reactions. They may also be used for preparing other hydrides by reactions such as



where  $M = \text{Zn, Cd, Be, or Mg}$  and  $R = \text{CH}_3$  or  $\text{C}_2\text{H}_5$ .

Because of its large size, the electron cloud of  $\text{H}^-$  is easily polarizable, and therefore,  $\text{H}^-$  is a soft base. Consequently, complexes with transition metals are usually formed in which the metals are soft Lewis acids. The metals are usually in low oxidation states, which makes them softer acids than the same metals in higher oxidation states. Typical among these compounds are  $\text{Fe}(\text{CO})_4\text{H}_2$ ,  $\text{Re}(\text{CO})(\text{P}(\text{C}_6\text{H}_5)_3)_3\text{H}$ , and  $\text{Mn}(\text{CO})_5\text{H}$ .

### 7.3.2 Interstitial Hydrides

Unlike the ionic hydrides, interstitial hydrides may have positive heats of formation, and they always have densities that are lower than the parent metal. Interstitial hydrides are generally described as the larger spherical atoms of the metal forced slightly apart with a smaller hydrogen atom occupying (interstitial) holes between them. Metallic palladium can absorb up to 700 times its own volume of hydrogen gas. Frequently, these compounds are characterized by formulas such as  $\text{CuH}_{0.96}$ ,  $\text{LaH}_{2.78}$ ,  $\text{TiH}_{1.21}$ ,  $\text{TiH}_{1.7}$ , or  $\text{PdH}_{0.62}$ .

In the formation of interstitial hydrides, the metal lattice is expanded to accommodate the hydrogen atoms, but the metal atoms retain their original crystal structure. There is, however, a significant change in physical and chemical properties of the metal. These properties are predictable in a very straightforward manner.

In a metal, certainly the transition metals, the electrons are more or less free to move in conduction bands. This fact is responsible for the high electrical conductivity of metals. When hydrogen atoms are present in the holes between the atoms, the movement of the electrons is somewhat impaired. As a result, the metal hydrides of this class are poorer conductors than the pure metals. The presence of hydrogen atoms makes the metal atoms less mobile and more restricted to particular lattice sites. Accordingly, the interstitial metal hydrides are more brittle than the parent metal. Also, the inclusion of the hydrogen atoms causes a small degree of lattice expansion so that the interstitial hydrides are less dense than the parent metal alone.

The nature of the process of forming the interstitial metal hydrides explains why some of these compounds are accompanied by a positive heat of formation. The bond energy in  $\text{H}_2$  is about  $432.6 \text{ kJ mol}^{-1}$ , and the metal lattice must be expanded to allow hydrogen atoms to enter the lattice. The interaction between the hydrogen atoms and metal atoms is not exothermic enough to compensate for the large energy requirement. The transition metals have electronegativities that are higher than the alkali or alkaline earth metals (generally in the range 1.4–1.7), which makes it unlikely that the hydrides of these metals would be ionic. Although the transition metals, lanthanides, and actinides have holes in their lattices that can accommodate hydrogen atoms, these hydrides probably should also not be considered as covalent. In some ways, the interstitial hydrides are probably best considered as solutions of atomic hydrogen in the metals. Because these metals interact with molecular hydrogen causing it to be separated into atoms, these metals can function as effective hydrogenation catalysts. These metals essentially function to split the hydrogen molecules into atoms, which, for the most part, remain in the metal lattice. Some escape of hydrogen does occur, and the hydrogen is evolved as highly reactive atomic hydrogen at the surface of the metal.

The prior treatment that the metal surface has received greatly affects the ease of hydride formation because the hydrogen must first be adsorbed on the metal surface before dissolution occurs. As a result, it is possible that not all of the available interstitial positions will become occupied by hydrogen resulting in compositions that are variable depending on the temperature and pressure used in preparing the metal hydride. As a result, the composition may not be exactly stoichiometric, and hydrides of this type are sometimes referred to as *nonstoichiometric hydrides*.

### 7.3.3 Covalent Hydrides

Strictly speaking, a binary compound of hydrogen is not a *hydride* unless hydrogen has the higher electronegativity of the two elements. Obviously, this is the case in the metal compounds such as the ionic and interstitial hydrides just discussed. There are, however, some covalent compounds in which this is also the case depending on the electronegativity of the other element. To illustrate this point, [Table 7.5](#) shows the electronegativities of some main group elements.

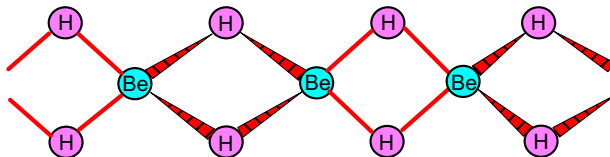
On the basis of electronegativities, we would say that  $\text{SbH}_3$  is a hydride whereas  $\text{NH}_3$  is not. Although any binary compound of hydrogen is sometimes referred to as a “hydride,” we will not use that classification here. Thus,  $\text{H}_2\text{O}$  is an *oxide* of hydrogen rather than a *hydride* of oxygen. Similarly,  $\text{HF}$ ,  $\text{HCl}$ ,  $\text{HBr}$ , and  $\text{HI}$  are considered to be hydrogen

**TABLE 7.5** Electronegativities of Some Main Group Elements

Element	Electronegativity	Element	Electronegativity
H	2.1	P	2.1
B	2.0	As	2.0
C	2.5	Sb	1.8
Si	1.8	F	4.0
Ge	1.8	Cl	3.0
N	3.0	Br	2.8

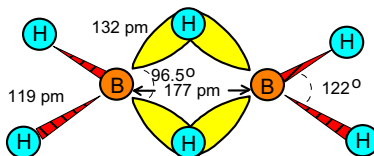
halides, not halogen hydrides. Typically, the binary compounds of B, Si, Ge, Sn, P, As, and Sb with hydrogen are considered to be hydrides.

Beryllium and magnesium hydrides,  $\text{BeH}_2$  and  $\text{MgH}_2$ , appear to be polymeric covalent hydrides rather than ionic hydrides as are those formed by the other Group II metals. The structure of these compounds consists of chains in which the H atoms form bridges as illustrated below for  $\text{BeH}_2$ .



The four H atoms around each Be are arranged in an approximately tetrahedral manner.

When magnesium is heated with boron, the product is magnesium boride,  $\text{MgB}_2$  (sometimes written as  $\text{Mg}_3\text{B}_2$ ). The hydrolysis of the compound produces  $\text{Mg}(\text{OH})_2$  and a boron hydride. Instead of the expected product of  $\text{BH}_3$ , borane, the product is  $\text{B}_2\text{H}_6$ , diborane. This compound has the structure



in which the boron atoms are located in the centers of somewhat distorted tetrahedra of hydrogen atoms (see Chapter 9). Diborane can also be prepared by the reaction of  $\text{BF}_3$  with  $\text{NaBH}_4$ .



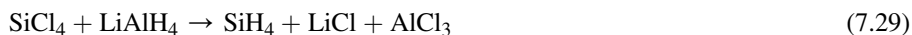
Many other boron hydrides are known, and the chemistry of these interesting compounds will be discussed in Chapter 9.

Because most of the covalent hydrides consist of molecular units with only weak intermolecular forces between them, they are volatile compounds. Accordingly, the covalent hydrides are sometimes referred to as the *volatile hydrides*. The nomenclature, melting points, and boiling points of several covalent hydrides are shown in Table 7.6.

Procedures similar to those described for the preparation of diborane can be employed for the preparation of the hydrides of silicon and germanium. For example,  $\text{Mg}_2\text{Si}$  results when Mg and Si react. This compound reacts with HCl to produce silane,  $\text{SiH}_4$ .



Lithium aluminum hydride reacts with  $\text{SiCl}_4$  as follows to form silane.



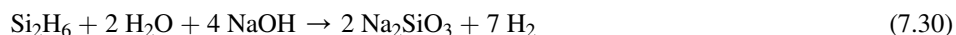
The reaction is not actually this simple, and a mixture of silanes is produced.



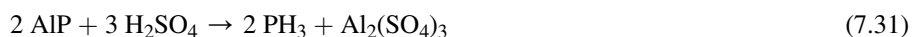
**TABLE 7.6** Properties of Some Covalent Hydrides

Name	Formula	m.p., °C	b.p., °C
Diborane	B <sub>2</sub> H <sub>6</sub>	-165.5	-92.5
Tetraborane	B <sub>4</sub> H <sub>10</sub>	-120	18
Pentaborane-9	B <sub>5</sub> H <sub>9</sub>	-46.6	48
Pentaborane-11	B <sub>5</sub> H <sub>11</sub>	-123	63
Hexaborane	B <sub>6</sub> H <sub>10</sub>	-65	110
Ennaborane	B <sub>9</sub> H <sub>15</sub>	2.6	—
Decaborane	B <sub>10</sub> H <sub>14</sub>	99.7	213
Silane	SiH <sub>4</sub>	-185	-119.9
Disilane	Si <sub>2</sub> H <sub>6</sub>	-132.5	-14.5
Trisilane	Si <sub>3</sub> H <sub>8</sub>	-117	54
Germane	GeH <sub>4</sub>	-165	-90
Digermane	Ge <sub>2</sub> H <sub>6</sub>	-109	29
Trigermane	Ge <sub>3</sub> H <sub>8</sub>	-106	110
Phosphine	PH <sub>3</sub>	-133	-87.7
Diphosphine	P <sub>2</sub> H <sub>4</sub>	-99	-51.7
Arsine	AsH <sub>3</sub>	-116.3	-62.4
Stibine	SbH <sub>3</sub>	-88	-18

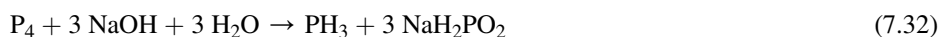
Silanes are more reactive than the corresponding carbon compounds, and silane is flammable in air. Although ethane is unreactive toward water, silanes such as Si<sub>2</sub>H<sub>6</sub> undergo hydrolysis in basic solution.



Acid hydrolysis of a metal phosphide produces phosphine.



The action of sodium hydroxide on elemental white phosphorus also produces phosphine.



This reaction also produces small amounts of diphosphine, P<sub>2</sub>H<sub>4</sub>, which is spontaneously flammable in air. Phosphine also burns readily in air,



Although PH<sub>3</sub> is somewhat similar to ammonia, it is a much weaker base toward protons in accord with the hard–soft interaction principle (see Chapter 6). Phosphonium salts are usually stable only when the anion is also large (e.g., PH<sub>4</sub>I) and the parent acid is a strong one (e.g., HBr or HI). However, PH<sub>3</sub> and substituted phosphines are very good ligands toward second and third row transition metal ions. Many coordination compounds containing such ligands are known, and some of them have important catalytic properties.

The bonding in PH<sub>3</sub> apparently involves much more nearly *s*–*p* overlap than is present in NH<sub>3</sub>. In NH<sub>3</sub>, the H–N–H bond angles are about 107° whereas in PH<sub>3</sub> the H–P–H angles are about 93°. Although increasing the extent of hybridization reduces repulsion by giving bond angles closer to the tetrahedral angle, it decreases the effectiveness of the overlap of the small hydrogen 1s orbital with the *sp*<sup>3</sup> orbitals on the central atom. The bond angles in AsH<sub>3</sub> and SbH<sub>3</sub> are even slightly smaller than those in PH<sub>3</sub>. A similar trend is seen for the series H<sub>2</sub>O, H<sub>2</sub>S, H<sub>2</sub>Se, and H<sub>2</sub>Te. In the hydrogen compounds of the heavier members of each group, there is very little tendency for the central atom to form *sp*<sup>3</sup> hybrids.

Thus, the H–X–H bond angles indicate that the central atom uses essentially pure  $p$  orbitals in bonding to the hydrogen atoms, and the bond angles in these compounds are approximately  $90^\circ$ .

The discussion of the covalent hydrides given here is somewhat brief in keeping with the intended purpose of this book. Some of the types of hydrides will be discussed in greater detail in chapters dealing with the other elements. The discussion presented here should serve to introduce the breadth, scope, and importance of hydride chemistry.

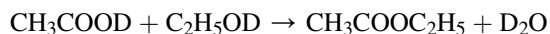
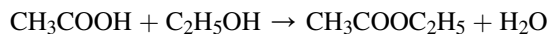
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## PROBLEMS

- Explain why there are more covalent than ionic binary compounds of hydrogen.
- Write balanced equations for the following processes.
  - Dehydrogenation of pentane
  - The reaction of carbon with high-temperature steam
  - The reaction of sodium hydride with water
  - The electrolysis of water
  - The reaction of zinc with aqueous HCl
- Complete and balance the following.
  - $\text{Li} + \text{H}_2 \rightarrow$
  - $\text{Ca}_3\text{N}_2 + \text{H}_2 \rightarrow$
  - $\text{CaH}_2 + \text{CH}_3\text{OH} \rightarrow$
  - $\text{Zn} + \text{NaOH} + \text{H}_2\text{O} \rightarrow$
  - $\text{Cd}(\text{C}_2\text{H}_5) + \text{LiAlH}_4 \rightarrow$
- Describe the steam reformer process for producing hydrogen.
- Describe some of the major difficulties associated with a hydrogen gas-based energy program.
- Write the equations to describe the preparation of the following compounds. Give conditions where possible.
  - $\text{H}_2$  (commercial production)
  - Diborane starting with magnesium metal and boron
  - NaOH (commercial production)
  - Styrene,  $\text{C}_6\text{H}_5\text{CH}=\text{CH}_2$
- Given the stability of the water molecule, speculate on the type of processes that could be used to produce the large quantities of hydrogen necessary for the gas to serve as the major energy source. Keep in mind that the processes must produce the gas cheaply enough for it to be economically competitive with other energy sources.
- Complete and balance the following.
  - $\text{LiAlH}_4 + \text{SnCl}_4 \rightarrow$
  - $\text{Ca} + \text{H}_2\text{O} \rightarrow$
  - $\text{AlCl}_3 + \text{LiH} \rightarrow$
  - $\text{CaH}_2 + \text{H}_2\text{O} \rightarrow$
  - $\text{Al} + \text{NaOH} + \text{H}_2\text{O} \rightarrow$
- Using the hard–soft interaction principle, explain why the apparent ionic radius of  $\text{H}^-$  in LiH is smaller than it is in KH.

10. Which of the following reactions would take place more rapidly if all other conditions are the same? Provide an explanation for your answer.



11. How does the electrolysis of water lead to the production of  $\text{D}_2\text{O}$ ?
12. Explain why acetylene,  $\text{C}_2\text{H}_2$ , behaves as an acid and reacts with active metals to liberate hydrogen.
13. Explain why a linear structure for  $\text{H}_3^+$  would be less stable than a triangular one.
14. From the standpoint of structure and bonding, explain why transition metals such as palladium and nickel are often used as catalysts for hydrogenation reactions.
15. Explain the following observations.
- (a) Interstitial hydrides are less dense than the parent metal.
  - (b) Interstitial hydrides have variable composition.
  - (c) Interstitial hydrides are poorer conductors of electricity than the parent metal.
  - (d) Ionic hydrides are more dense than the parent metal.

## Chapter 8

# The Group IA and IIA Metals

The elements that constitute Groups IA and IIA of the periodic table are the active metals known as the alkali metals and alkaline earths, respectively. Because of their reactivity, none of these elements is found free in nature. Moreover, because they are difficult to reduce,



they were not easy enough to reduce by the chemical means available to the ancients. Therefore, the uncombined elements were not obtained until comparatively recently. Most of the metals were obtained by electrochemical means in the early 1800s. However, several compounds of the Group IA and IIA metals have been known since the earliest times. Salt (NaCl), limestone (CaCO<sub>3</sub>), and sodium carbonate (Na<sub>2</sub>CO<sub>3</sub>) have had important uses throughout history.

The names of the elements in Groups IA and IIA are derived from words in other languages. For example, the name lithium comes from the Greek *lithos* meaning stone. The symbols Na for sodium and K for potassium come from the Latin names *natrium* and *kalium*, respectively. The name rubidium comes from the Latin word *rubidius* meaning deepest red. The name cesium comes from the Latin word *caesius*, which means sky blue.

Elements in Group IIA have names derived from words in other languages also. For example, beryllium comes from the Greek word *beryllos* for beryl, the mineral containing beryllium. Magnesium is derived from the name for the Magnesia district in Greece. The name calcium is derived from the Latin word *calx* meaning lime. Strontium is named after Strontian, Scotland because it was first recognized as an element in 1790 by A. Crawford in Edinburgh. It was isolated by Sir Humphry Davy in 1808. The name barium is derived from the Greek word *barys* meaning heavy, whereas the Latin word *radius*, meaning ray, is the source of the name radium.

Beryllium was first prepared by F. Whöler in 1828, the same year in which the distinction between inorganic and organic chemistry disappeared when he converted ammonium cyanate into urea. The reaction utilized by Whöler was the reduction of BeCl<sub>2</sub> with potassium,



It was prepared independently about the same time by A. A. B. Bussy.

## 8.1 GENERAL CHARACTERISTICS

The elements in Group IA and IIA of the periodic table have the valence shell configurations of  $ns^1$  and  $ns^2$ , respectively. The elements in Group IA have the lowest ionization potentials of any group in the periodic table. It is not unexpected that the most common characteristic of Group IA chemical behavior is their tendency to form +1 ions. All naturally occurring compounds of the Group IA elements contain the elements in that form. [Tables 8.1 and 8.2](#) show information on the occurrence of Group IA elements, their physical properties, and some major uses of the elements and their compounds.

In addition to the data shown in [Tables 8.1 and 8.2](#), other properties are important in studying systematically the chemistry of the Group IA elements. The data shown in [Table 8.3](#) give some of the necessary information. The electrostatic attraction between charged species (either ions or ion dipole) depends on the charges and sizes on the species. Although the charge/volume ratio is used in some cases, the values listed here are simply the charge divided by the radius in pm. This approach is adequate in most cases.

Compounds of the alkaline earth metals occur widely in nature, and some of them have been known since antiquity. Calcium carbonate, of which limestone is one form, has been used as a building material since the Stone Age and as a source of lime (CaO) for thousands of years. Calcium and magnesium are among the elements of greatest abundance. [Tables 8.4 and 8.5](#) show some of the information on occurrence, production, properties, and uses of the Group IIA metals.

**TABLE 8.1** Composition and Sources of the Group IA Minerals

	Minerals	Composition	Mineral Sources <sup>a</sup>
Li	Spodumene	LiAlSi <sub>2</sub> O <sub>6</sub>	U.S., Canada, Africa, Brazil, Argentina
	Amblygonite	(Li,Na)AlPO <sub>4</sub> (F,OH)	
	Lepidolite	K <sub>2</sub> Li <sub>3</sub> Al <sub>4</sub> Si <sub>7</sub> O <sub>21</sub> (OH,F) <sub>3</sub>	
Na	Salt	NaCl	Many places in the world
	Trona	Na <sub>2</sub> CO <sub>3</sub> ·NaHCO <sub>3</sub> ·2 H <sub>2</sub> O	
K	Carnallite	KMgCl <sub>3</sub> ·6 H <sub>2</sub> O	Germany, Canada, Great Salt Lake, England, Israel, Russia
	Sylvite	KCl	
	Polyhalite	K <sub>2</sub> Ca <sub>2</sub> Mg(SO <sub>4</sub> ) <sub>4</sub> ·2 H <sub>2</sub> O	
Rb	Carnallite	KMgCl <sub>3</sub> ·6 H <sub>2</sub> O	Many places in the world in salt brines and mineral water
	Pollucite	CsAlSi <sub>2</sub> O <sub>6</sub>	
Cs	Pollucite	CsAlSi <sub>2</sub> O <sub>6</sub>	Canada, S. Africa, U.S. (ME, SD)

<sup>a</sup>Not all of the minerals occur in each country.

**TABLE 8.2** Properties and Uses of the Group IA Metals

	Production	Crystal Structure <sup>a</sup>	Density (g cm <sup>-3</sup> )	m.p., °C	b.p., °C	Uses
Li	Electrolysis of molten LiCl + KCl	<i>bcc</i>	0.534	180.5	1342	Li grease, soap, batteries, LiAlH <sub>4</sub>
Na	Electrolysis of molten NaCl + CaCl <sub>2</sub>	<i>bcc</i>	0.970	97.8	883	Reducing agent, lamps, nuclear reactors
K	Reaction of Na with KCl at high temperature	<i>bcc</i>	0.862	63.3	760	Alloys with Na, miscellaneous chemicals
Rb	Reaction of Ca with RbCl at high temperature, electrolysis of molten RbCl	<i>bcc</i>	1.53	39	686	Catalysts, photocells
Cs	Reaction of Ca with CsCl at high temperature, electrolysis of molten CsCN	<i>bcc</i>	1.87	28	669	Catalysts, photocells
Fr	Longest lived isotope has a half life of 21 min					

<sup>a</sup>The body centered cubic structure is indicated by *bcc*.

**TABLE 8.3** Characteristics of Atoms and Ions of Group IA Elements

Metal	Radii (pm), Atom	Radii (pm), Ion	M–M Bond Energy (kJ mol <sup>-1</sup> )	Ionization Enthalpy (kJ mol <sup>-1</sup> )	–Δ <i>H</i> <sub>hyd</sub> <sup>a</sup> (kJ mol <sup>-1</sup> )	Charge to Size Ratio
Li	152	68	108.0	520.1	515	0.015
Na	186	95	73.3	495.7	406	0.011
K	227	133	49.9	418.7	322	0.0075
Rb	248	148	47.3	402.9	293	0.0068
Cs	265	169	43.6	375.6	264	0.0059

<sup>a</sup>The negative value of the enthalpy of hydration of the gaseous ions.

**TABLE 8.4** Composition and Sources of the Alkaline Earth Metals

	Minerals	Composition	Mineral Sources <sup>a</sup>
Be	Beryl	Be <sub>3</sub> Al <sub>2</sub> Si <sub>6</sub> O <sub>18</sub>	S. Africa, U.S. (CO, ME, NH, SD)
Mg	Magnesite	MgCO <sub>3</sub>	Many places throughout the world, sea water and brines
	Dolomite	CaMg(CO <sub>3</sub> ) <sub>2</sub>	
Ca	Calcite	CaCO <sub>3</sub>	Many places throughout the world
	Dolomite	CaMg(CO <sub>3</sub> ) <sub>2</sub>	
	Gypsum	CaSO <sub>4</sub> ·2 H <sub>2</sub> O	
Sr	Strontianite	SrCO <sub>3</sub>	Germany, England, U.S. (NY, PA, CA)
Ba	Barite	BaSO <sub>4</sub>	U.S. (GA, MO, AR, KY, CA, NV), Canada, Mexico
	Witherite	BaCO <sub>3</sub>	
Ra	Pitchblende	UO <sub>2</sub> ore	Canada, Zaire, France, Russia, U.S. (CO)
	Carnotite	K <sub>2</sub> (UO <sub>2</sub> ) <sub>2</sub> (VO <sub>4</sub> ) <sub>2</sub> ·3 H <sub>2</sub> O	

<sup>a</sup>Not all of the minerals occur in each country.

**TABLE 8.5** Properties and Uses of the Group IIA Metals

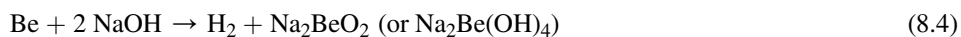
	Production	Crystal Structure <sup>a</sup>	Density (g cm <sup>-3</sup> )	m.p., °C	b.p., °C	Uses
Be	Oxide to halide then reduce w/Mg or electrolysis	<i>hcp</i>	1.85	1278	2970	Neutron moderator, X-ray tubes, electrical equipment
Mg	MgCl <sub>2</sub> electrolysis or MgO + ferrosilicon	<i>hcp</i>	1.74	649	1090	Alloys for cars, reducing agent, aircraft, organics
Ca	CaO + Al (at high temperature) or electrolysis of molten CaCl <sub>2</sub>	<i>fcc</i>	1.54	839	1484	Alloys, reducing agent
Sr	SrO + Al (at high temperature) or electrolysis of molten SrCl <sub>2</sub>	<i>fcc</i>	2.54	769	1384	Alloys, fireworks, specialty glasses, sugar refining
Ba	BaO + Al (at high temperature in a vacuum)	<i>bcc</i>	3.51	725	1640	Alloys
Ra	Electrolysis of Ra salts	—	5	700	1140	Radiography medical uses

<sup>a</sup>The structures indicated are as follows: *fcc* = face centered cubic; *bcc* = body centered cubic; *hcp* = hexagonal close packing.

As will be discussed in greater detail later, some of the naturally occurring minerals of the Group IIA metals are of great importance. Some of the properties of atoms and ions of the Group IIA elements are shown in Table 8.6.

Chemically, the free metals of Group IA and IIA are all reducing agents, so they are capable of displacing other metals from compounds. Because of their reactivity, they will react with most nonmetals to form binary compounds. Beryllium is used in alloys with copper. The strength of copper is greatly increased by adding 1–2% beryllium, and these alloys are widely used in the fabrication of objects that must have good electrical conductivity and wear resistance.

Beryllium compounds are substantially more covalent than are those of the other Group IIA metals as a result of the high charge to size ratio that results from a +2 ion having a radius of only about 30 pm. From a comparison of the properties of the Group IIA elements it can be seen that beryllium is considerably different from other members of the group. It is also amphoteric and like aluminum dissolves in strong bases to liberate hydrogen.



**TABLE 8.6** Characteristics of Atoms and Ions of Group IIA Elements

Metal	Radii (pm), Atom	Radii (pm), Ion	Ionization Enthalpy		$-\Delta H_{\text{hyd}}^a$ (kJ mol <sup>-1</sup> )	Charge to Size Ratio
			First (kJ mol <sup>-1</sup> )	Second (kJ mol <sup>-1</sup> )		
Be	111	31	899	1757	1435	0.065
Mg	160	65	737	1450	2003	0.031
Ca	197	99	590	1146	1657	0.020
Sr	215	113	549	1064	1524	0.018
Ba	217	135	503	965	1360	0.015
Ra	220	140	509	979	~1300	0.014

<sup>a</sup>The negative value of the heat of hydration of the gaseous ions.

It should be recognized that the  $\text{Na}_2\text{BeO}_2$  and  $\text{Na}_2\text{Be}(\text{OH})_4$  are equivalent if the latter has two molecules of water removed. Also as a result of the small size of the  $\text{Be}^{2+}$  ion, beryllium is four coordinate in most of its compounds but the other members of Group IIA are most often six coordinate.

One of the interesting characteristics of the metals in Groups IA and IIA is their solubility in liquid ammonia with the Group IA metals being more soluble than those of Group IIA. Some of the chemical and physical aspects of these solutions have been discussed in Chapter 6.

In general, the reactivity of the Group IA metals increases in progressing down the group. This is due in part to the decrease in ionization enthalpy in going down the group. Thus, one would expect that the formation of cesium compounds would be more exothermic than the formation of those of rubidium. By the same argument, the formation of the rubidium compounds should be more energetically favorable than those of potassium, those of potassium more favorable than those of sodium, and those of sodium more favorable than those of lithium. This line of reasoning is in fact correct for the chlorides, bromides, and iodides. However, the reverse order of stability is seen for the alkali fluorides where  $\text{LiF}$  is the most stable and  $\text{CsF}$  the least stable.

The ionization of a gaseous atom is simpler than the formation of a compound where a solid lattice is produced. Because of the small size of  $\text{Li}^+$ , the lattice energies of the lithium compounds with a given anion are greater than those in which the other alkali metal ions are present. As a result, the lithium salts follow the expected trend in stability predicted above except for  $\text{LiF}$  where the large lattice energy results from the interaction of small positive and negative ions. It is a general trend that the compounds of  $\text{Li}$  with small anions ( $\text{F}^-$ ,  $\text{O}^{2-}$ ,  $\text{N}^{3-}$ , etc.) are the most stable of the Group IA compounds with these anions. For example,  $\text{Li}$  is the only Group IA metal that gives a regular oxide (see Section 8.2), and it reacts more easily with  $\text{N}_2$  than do the remainder of the elements in the group.

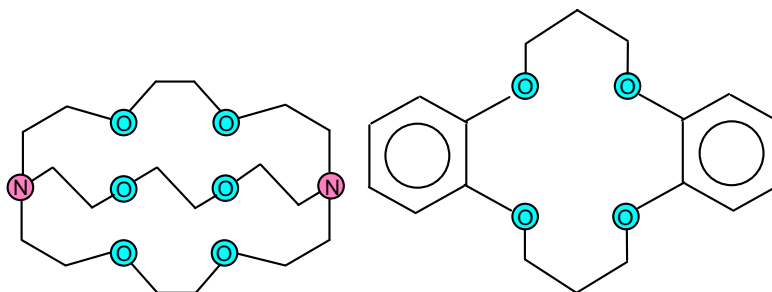
Removing an electron from a sodium atom requires  $496 \text{ kJ mol}^{-1}$  although adding an electron to another sodium atom releases approximately  $53 \text{ kJ mol}^{-1}$ . Thus, the process



would have an enthalpy change of  $+443 \text{ kJ mol}^{-1}$ . The data given in Table 8.3 show that the process



liberates  $406 \text{ kJ mol}^{-1}$ . Of course sodium reacts vigorously with water, but if a different solvent could be used that strongly solvated  $\text{Na}^+(g)$  the solvation enthalpy might overcome the  $+443 \text{ kJ mol}^{-1}$  required for electron transfer. The  $\text{Na}^-(g)$  would also be solvated, but its solvation enthalpy would be less than that of the smaller  $\text{Na}^+(g)$ . Such a situation has been devised in which sodium is dissolved in ethylenediamine (the solubility of alkali metals in liquid ammonia was discussed in Chapter 6). In this case, a complexing agent,  $\text{N}[\text{CH}_2\text{CH}_2\text{OCH}_2\text{CH}_2\text{OCH}_2\text{CH}_2]_3\text{N}$  (see Figure 8.1) is referred to as 2,2,2-cryptand or simply 2,2,2-crypt. It binds strongly to  $\text{Na}^+$  so that the  $\text{Na}^+$  that forms is complexed. When the ethylenediamine is evaporated, a solid compound is formed, which contains the strongly bonded  $\text{Na}^+$  cation and the  $\text{Na}^-$  anion. The crown ether known as dibenzo-14-crown-4 is also shown in Figure 8.1, and it forms similar types of complexes.



**FIGURE 8.1** Two complexing agents that form stable complexes with alkali metal ions. The ethylenediamine derivative (left) (known as 2,2,2-cryptand or 2,2,2-crypt) contains eight potential bonding atoms, and the crown ether (right) contains four electron pair donor atoms.

The stability of complexes of alkali metals with various complexing agents is known to depend on matching the size of the alkali metal to the size of the opening in the molecule. However, for most complexes of the ions of Group IA metals, the stability decreases in the order  $\text{Li}^+ > \text{Na}^+ > \text{K}^+$  in accord with the decrease in the charge to size ratio.

Most of the Group IA and IIA metals react with hydrogen to form metal hydrides. For all of the metals in these two groups except Be and Mg, the hydrides are considered to be ionic or saltlike hydrides containing  $\text{H}^-$  ions (see Chapter 7). The hydrides of beryllium and magnesium have considerable covalent character. The molten ionic compounds conduct electricity as do molten mixtures of the hydrides in alkali halides, and during electrolysis of the hydrides, hydrogen is liberated at the anode as a result of the oxidation of  $\text{H}^-$ .



Ionic hydrides are characterized chemically by the basicity of the hydride ion as illustrated by the reactions



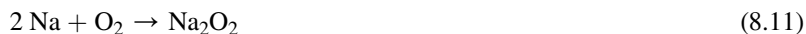
Because of the great affinity of  $\text{H}^-$  for protons, NaH and  $\text{CaH}_2$  are used as drying agents for removal of traces of water from organic solvents. Lithium aluminum hydride,  $\text{LiAlH}_4$ , is a strong reducing agent that is used extensively in reduction reactions in organic chemistry. The chemistry of ionic hydrides is presented in greater detail in Chapter 7.

## 8.2 OXIDES AND HYDROXIDES

Although all of the Group IA and IIA metals will react with oxygen, there is considerable variation in the ease of the reactions and the nature of the products. In Chapter 5, it was described how the reaction of oxygen with the Group IA metals produces a “normal” oxide only in the case of Li.



Sodium reacts with oxygen to produce a peroxide,



but potassium, rubidium, and cesium give the superoxides.



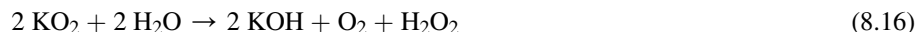
To some extent, the products can be mixtures of oxides with the composition depending on the reaction conditions.

At least part of the stabilization of the compounds containing the diatomic anions  $\text{O}_2^{2-}$  and  $\text{O}_2^-$  comes from the fact that crystal lattices are most stable when the cations and anions have similar sizes (see Chapter 6). The Group IIA metals give normal oxides except for barium and radium, which give peroxides when they react with oxygen.



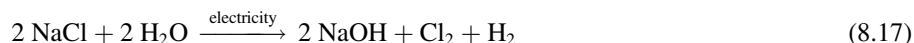


Oxides of metallic elements typically react with water to produce hydroxides. In the case of the oxides of the Group IA elements, the reactions are



The Group IA metal hydroxides are all strong bases with the base strength increasing in going from LiOH to CsOH for the gaseous compounds.

Sodium hydroxide is an extremely important base that is used to the extent of about 19 billion pounds annually. The compound is sometimes referred to as caustic soda or simply caustic. Today, it is made by the electrolysis of an aqueous solution of sodium chloride.

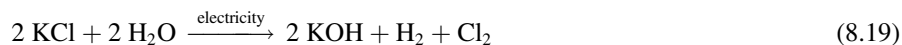


Not only does this process produce sodium hydroxide, but also chlorine and hydrogen. As a result, this is one of the most important processes in the chemical industry. Because of the competing reaction



the separation of the chlorine and sodium hydroxide formed must be effected. Two continuous processes are the use of a diaphragm cell and the use of a mercury cell. In the latter, a mercury cathode is used because sodium reacts readily with mercury to form an amalgam. The amalgam is continuously removed as it flows over the bottom of the cell, and after it is removed from the cell, the sodium is allowed to react with water to produce sodium hydroxide. In the diaphragm cell, a diaphragm made of asbestos separates the cathode and anode compartments. As the process takes place, the NaOH solution is allowed to flow out of the cathode compartment before it has a chance to diffuse through the asbestos filter into the anode compartment.

Potassium hydroxide, sometimes known as caustic potash, is produced in essentially the same way as sodium hydroxide.



It is used extensively in the manufacture of liquid soaps and detergents. Potassium hydroxide is more soluble in organic solvents, particularly alcohols, so strongly basic solutions in these solvents are frequently made to contain KOH rather than NaOH. Rubidium and cesium hydroxides are even stronger bases than NaOH and KOH, but they are of little practical importance because of cost and availability. Solutions of the alkali metal hydroxides often contain traces of the bicarbonates owing to the reaction of  $\text{CO}_2$ , an acidic oxide, with the hydroxide ions in the solutions.



The reactions of the oxides of Group IIA with water give the corresponding hydroxides.

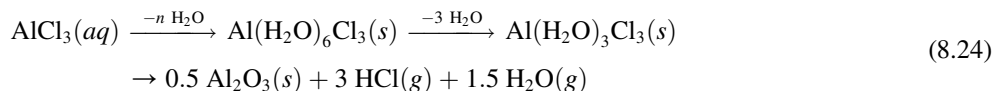


These compounds are bases except for beryllium oxide, which is amphoteric.

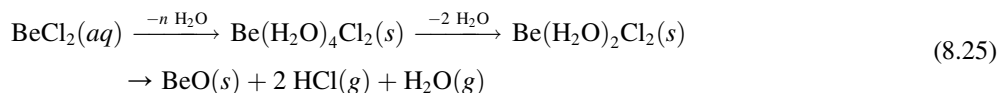


Magnesium hydroxide is a weak base that is used as an insoluble suspension known as “milk of magnesia.”

Beryllium does not resemble closely the other elements in Group IIA. The high heat of solvation of the small  $\text{Be}^{2+}$  ion and the greater degree of covalence of beryllium compounds result in their being generally more soluble than are the corresponding compounds of magnesium and calcium. The +2 beryllium ion has a radius of 31 pm whereas  $\text{Mg}^{2+}$  has a radius of 65 pm. The charge to size ratios for  $\text{Be}^{2+}$  and  $\text{Mg}^{2+}$  are 0.065 and 0.031, respectively, and the values for the +2 ions of other metals in the group are even smaller. The radius of  $\text{Al}^{3+}$  is 50 pm so the charge to size ratio is 0.060. As a result of their having similar charge to size ratios, the chemistry of  $\text{Be}^{2+}$  and  $\text{Al}^{3+}$  is in some ways quite similar. For example, if  $\text{AlCl}_3$  is dissolved in water, evaporation of the water does not yield recrystallized  $\text{AlCl}_3$ . As water is evaporated, intermediate stages having various degrees of hydration are produced, but continued heating results in the loss of HCl. The final product is aluminum oxide, not aluminum chloride. This behavior can be shown by a series of equations.



Beryllium shows the same affinity for forming bonds to oxygen, which prevents  $\text{Be}(\text{H}_2\text{O})_4\text{Cl}_2(s)$  from forming anhydrous  $\text{BeCl}_2$  by dehydration.



The fact that  $\text{Be}^{2+}$  resembles  $\text{Al}^{3+}$  is known as a *diagonal relationship* because Al is one column to the right and one row below Be along a diagonal in the periodic table.

The fact that  $\text{Be}^{2+}$  and  $\text{Al}^{3+}$  have charge to size ratios that are so high causes them to exert a great polarizing effect on the molecules and ions to which they are bound. This results in their compounds being substantially covalent. Unlike the situation when NaCl is dissolved in water and the ions become hydrated, dissolving  $\text{BeCl}_2$  or  $\text{AlCl}_3$  in water results in the formation of essentially covalent complexes such as  $\text{Be}(\text{H}_2\text{O})_4^{2+}$  and  $\text{Al}(\text{H}_2\text{O})_6^{3+}$  from which the water cannot easily be removed. When heated strongly, the loss of HCl becomes the energetically favorable process rather than loss of water from compounds such as  $\text{Be}(\text{H}_2\text{O})_4\text{Cl}_2$  and  $\text{Al}(\text{H}_2\text{O})_6\text{Cl}_3$  because of the strong bonds to oxygen formed by  $\text{Be}^{2+}$  and  $\text{Al}^{3+}$ .

The oxides of the other Group IIA metals are strong bases although they have very limited solubility in water. For example,  $\text{Ca}(\text{OH})_2$  dissolves to the extent of about 0.12 g in 100 g of water.  $\text{Ca}(\text{OH})_2$  is a strong base that is obtained by heating limestone to produce lime,



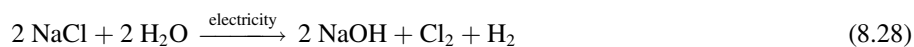
and the lime reacts with water to produce hydrated or slaked lime,  $\text{Ca}(\text{OH})_2$ .



Calcium hydroxide is an extremely important base (34 billion pounds of the oxide and hydroxide are produced annually) that is used on a very large scale because it is less expensive than sodium or potassium hydroxide. One of the major uses of  $\text{Ca}(\text{OH})_2$  is in the production of mortar. In this use, the  $\text{Ca}(\text{OH})_2$  reacts with carbon dioxide from the atmosphere to produce  $\text{CaCO}_3$  (limestone), which binds the particles of sand and gravel together in concrete.

### 8.3 HALIDES

Sodium chloride is found in enormous quantities throughout the world, and it is one of the most useful naturally occurring inorganic chemicals. It is mined in several places in the world and it occurs in salt beds, salt brines, sea water, and other sources. Many of the important sodium compounds are produced with sodium chloride as the starting material. About 50% of the NaCl is consumed in the processes that are carried out to produce sodium hydroxide and chlorine. The process illustrated in Eq. (8.28) is the source of most of the 19 billion pounds of NaOH and the 22 billion pounds of chlorine produced annually.



A considerable amount is also used in the process to produce sodium and chlorine by the electrolysis of molten NaCl.



Because of the high melting point of sodium chloride (801 °C), the electrolysis of a lower melting eutectic with  $\text{CaCl}_2$  is also carried out. Large amounts of NaCl are used in food industries and to melt snow and ice from highways and sidewalks. For the latter,  $\text{CaCl}_2$  is more effective owing to the fact that an aqueous mixture containing approximately 30%  $\text{CaCl}_2$  melts at  $-50$  °C but the lowest melting mixture of NaCl and water melts at  $-18$  °C.

Historically, the Solvay process for preparing sodium carbonate required a large amount of NaCl for the overall reaction



After separating the solid  $\text{NaHCO}_3$ , heating it gives  $\text{Na}_2\text{CO}_3$ .



The Solvay process is still important on a global basis, but in the United States,  $\text{Na}_2\text{CO}_3$  is obtained from the mineral *trona*,  $\text{Na}_2\text{CO}_3 \cdot \text{NaHCO}_3 \cdot 2 \text{H}_2\text{O}$ . Over 23 billion pounds of  $\text{Na}_2\text{CO}_3$  are produced annually from trona.

Beryllium halides are Lewis acids that form adducts with many types of electron pair donors (amines, ethers, phosphines, etc.) and they readily form complex ions. For example the  $\text{BeF}_4^{2-}$  ion results from the reaction



In these cases where there are four bonds to beryllium the bonding is approximately tetrahedral around the central atom.

In the vapor phase, dihalides of the Group IIA metals are expected to have a linear structure. However, the bond angles in  $\text{CaF}_2$ ,  $\text{SrF}_2$ , and  $\text{BaF}_2$  are estimated to be approximately  $145^\circ$ ,  $120^\circ$ , and  $108^\circ$ , respectively. Although linear structures are predicted on the basis of the valence shell electron pair repulsion (VSEPR) model, the energy required to bend these molecules is low enough that subtle factors result in bent structures. It is believed that two factors that may be involved are participation of the inner *d* orbitals and the polarization of these orbitals. Although the theory of bonding in these compounds will not be described in detail, these examples serve to show that simple approaches to chemical bonding are not always sufficient to explain the structures of some relatively simple molecules.

## 8.4 SULFIDES

All of the Group IA and IIA metals form sulfides, some of which are used rather extensively. The sulfides of Group IIA metals consist of  $\text{M}^{2+}$  and  $\text{S}^{2-}$  ions arranged in the sodium chloride type lattice (see Chapter 4). The compounds of the Group IA metals consist of  $\text{M}^+$  and  $\text{S}^{2-}$ , but as a result of there being twice as many cations as anions, the structure is of the antifluorite type (see Chapter 4). The sulfide ion is a base so there is extensive hydrolysis in solutions of the sulfides, and the solutions are basic.

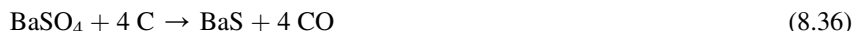


Sulfides of the Group IA and IIA metals can be obtained by the reaction of  $\text{H}_2\text{S}$  with the metal hydroxides.



One of the most interesting characteristics of solutions of the metal sulfides is their ability to dissolve sulfur with the formation of polysulfide ions,  $\text{S}_n^{2-}$  (in most cases,  $n < 6$ ). Some solid compounds containing polysulfide ions can be isolated, especially with large cations (see Chapter 16).

The most important sulfides of the Group IA and IIA metals are  $\text{Na}_2\text{S}$  and  $\text{BaS}$ . They can be prepared by the reduction of the sulfates by heating them with carbon at very high temperature.



Sodium sulfide has been used in tanning leather and in the manufacture of dyes. *Lithopone* is a pigment that contains barium sulfate and zinc sulfide that is made by the reaction



Calcium sulfide can be prepared by the reaction of  $\text{H}_2\text{S}$  with hydrated lime,  $\text{Ca}(\text{OH})_2$ ,



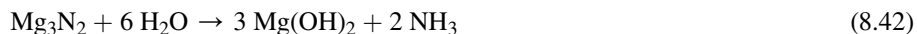
## 8.5 NITRIDES AND PHOSPHIDES

The alkali metals and alkaline earth metals are reactive toward most nonmetallic elements. Accordingly, they react with nitrogen, phosphorus, and arsenic to give binary compounds, but only those of nitrogen and phosphorus will be considered here. The reactions occur when the two elements are heated together as shown in the following equations.



Although the products are written as simple ionic binary compounds, it is known that complex materials containing anions that consist of polyhedral species containing the nonmetal are also produced. For example, a  $P_7^{3-}$  cluster is known, which has six of the phosphorus atoms arranged in a trigonal prism with the seventh occupying a position above the triangular face on one end of the prism.

The nitrides and phosphides of the Group IA and IIA metals contain anions of high charge, which behave as strong bases. Therefore, they abstract protons from a variety of proton donors. The following reactions are typical.



The nitrides and phosphides of the Group IA and IIA metals are not of great commercial importance.

## 8.6 CARBIDES, CYANIDES, CYANAMIDES, AND AMIDES

The reactive metals of Groups IA and IIA will react at elevated temperatures with carbon and silicon to give binary compounds. Of all the possible carbide compounds, by far the most important of the carbides is calcium carbide,  $\text{CaC}_2$ . This compound is properly considered as an acetylide because its reaction with water produces acetylene.



Other carbides, such as  $\text{Al}_4\text{C}_3$  should be considered as methanides because they react with water to produce methane.



Calcium acetylide is produced by heating lime and coke at very high temperature according to the equation

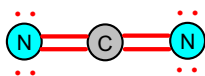


and it has had two important uses. First, it is used in the “carbide lamp” (not nearly as common as a generation ago) that was widely employed in mines. The lamp is constructed of two brass chambers with an upper chamber that contains water and the lower one that contains  $\text{CaC}_2$ . Water is allowed to drip through a small tube into the lower chamber where acetylene is generated. The gas escapes through a pinhole that is centered in a reflector. The gas jet is ignited by a flint and the lamp can be used for a considerable time before refilling is required.

A second important use of calcium acetylide is in the manufacture of calcium cyanamide,  $\text{CaCN}_2$ , which is prepared by the reaction of  $\text{CaC}_2$  with  $\text{N}_2$  at high temperature.



The cyanamide ion,  $\text{CN}_2^{2-}$ , contains 16 valence shell electrons and has the linear structure for which the principle resonance structure can be shown as



With the cyanamide ion having the arrangement of atoms  $\text{NCN}$ , calcium cyanamide can also be written as  $\text{CaNCN}$ . The production of calcium cyanamide is important because it represents one process leading to nitrogen fixation. Of course, the other is the production of ammonia by the Haber process (see Chapter 13).



Calcium cyanamide reacts with steam at high temperature to yield ammonia,



Calcium cyanamide has also been used extensively as a fertilizer for many years.

Sodium cyanamide is used mainly in the production of sodium cyanide that is used extensively in preparing solutions from which metals are electroplated. Sodium cyanide is also used in an extraction process for obtaining gold and silver

because these metals form stable complexes with  $\text{CN}^-$ . The sodium cyanamide is obtained by the reaction of sodium amide,  $\text{NaNH}_2$ , with carbon.



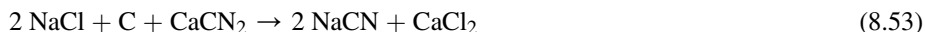
The sodium cyanamide reacts with carbon to produce the cyanide.



Sodium amide can be prepared by the reaction of liquid sodium with ammonia at  $400^\circ\text{C}$ .



Sodium cyanide can also be prepared from calcium cyanamide by the reactions



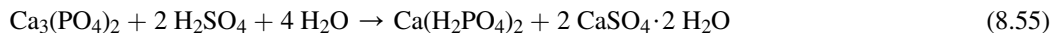
Sodium cyanide, an extremely toxic compound, is also used in the case hardening of steel objects (see Chapter 11).

## 8.7 CARBONATES, NITRATES, SULFATES, AND PHOSPHATES

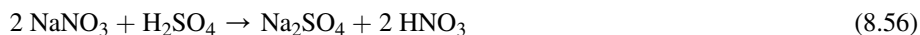
Some of the Group IA and IIA metals are found in nature in the form of carbonates, silicates, nitrates, and phosphates. For example, calcium carbonate is one of the most important naturally occurring compounds, and it is found in several forms. The most common form of calcium carbonate is limestone, which is used extensively as a building stone as well as the source of lime. Other forms include chalk, *calcite*, *aragonite*, Iceland spar, marble, and onyx. Many other materials such as egg shells, coral, pearls, and sea shells are composed predominantly of calcium carbonate. Thus, it is one of the most widely occurring compounds in nature.

Magnesium carbonate is found in nature as the mineral *magnesite* and in several other minerals. The silicate,  $\text{Mg}_2\text{SiO}_4$ , is found as the mineral *olivine*. Beryllium is found in the mineral *beryl*, which has the composition  $\text{Be}_3\text{Al}_2(\text{SiO}_3)_6$ . *Epsom salt*,  $\text{MgSO}_4 \cdot 7 \text{H}_2\text{O}$ , is a naturally occurring sulfate that is widely used to prepare a solution for soaking sprains, cuts, and other injuries. In addition, magnesium salts are found in sea water and this has been the commercial source of the metal. *Dolomite* has the composition  $\text{CaCO}_3 \cdot \text{MgCO}_3$  and it is also used as a building stone and in antacid preparations. Calcium sulfate is found as the mineral *gypsum*,  $\text{CaSO}_4 \cdot 2 \text{H}_2\text{O}$ .

Because of its use in the manufacture of fertilizers, calcium phosphate is a compound of enormous importance. After being mined, it is converted into  $\text{Ca}(\text{H}_2\text{PO}_4)_2$  by treating it with sulfuric acid. This converts the insoluble  $\text{Ca}_3(\text{PO}_4)_2$  into a soluble, more efficient form (see Section 14.9). The reaction can be written as



Sodium and potassium nitrates have been of enormous significance in the past because the preparation of nitric acid for many years was by means of the reaction



The major deposits of alkali nitrates found in Chile were vital for making nitric acid, which is necessary to prepare almost all types of explosives and propellants. Of course, nitric acid is now obtained by the catalytic oxidation of ammonia by the Ostwald process (see Chapter 13).

For many centuries, sodium carbonate has been recovered from the beds of dried lakes and inland seas. It is also found in the mineral *trona*,  $\text{Na}_2\text{CO}_3 \cdot \text{NaHCO}_3 \cdot 2 \text{H}_2\text{O}$ , which is now the commercial source of the compound in the United States. Sodium carbonate ranks high on the list of chemicals most used and the majority of it is used in the manufacture of glass.

From the discussion above, it should be apparent that the carbonates, sulfates, nitrates, and phosphates of the Group IA and IIA metals are of enormous importance in inorganic chemistry.

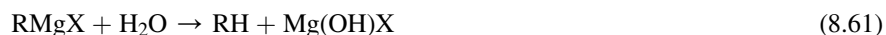
## 8.8 ORGANIC DERIVATIVES

Although Chapter 22 is devoted to the chemistry of organometallic compounds, a brief description of this type of compound of the metals in Groups IA and IIA is presented here. The most important of these metals with respect

to their organometallic compounds are lithium, magnesium, and sodium. Many organic derivatives of other metals in these groups are well known however. Although Zeise's salt, an organometallic complex having the formula  $K[Pt(C_2H_4)Cl_3]$ , was prepared earlier, the beginning of the organic chemistry of the metals in Groups IA and IIA can be regarded as the preparation of compounds such as  $C_2H_5MgBr$  by Victor Grignard in 1900 utilizing a reaction that can be shown as



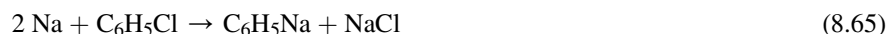
The organometallic compounds having the general formula  $RMgX$  are known as Grignard reagents, and they are of great utility in organic synthesis because they function as alkyl group transfer agents. Typical reactions of Grignard reagents are the following.



In addition to the magnesium compounds, lithium alkyls are versatile reagents that also function as alkyl group transfer agents. Lithium alkyls are prepared by the reaction of lithium with an alkyl halide using benzene or petroleum ether as a solvent.



A similar reaction to prepare phenyl sodium can be written as



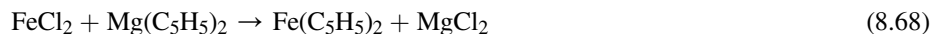
Cyclopentadiene forms alkali metal derivatives that are essentially ionic.



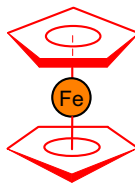
The magnesium compound is obtained by the analogous reaction that takes place at  $500^\circ C$  in a nitrogen atmosphere.



The ionic cyclopentadienyl compounds react with  $FeCl_2$  to give the "sandwich" compound known as ferrocene,  $Fe(C_5H_5)_2$ .

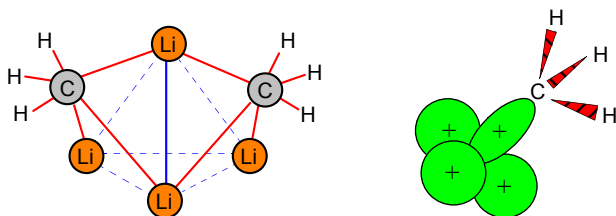


The structure of ferrocene is

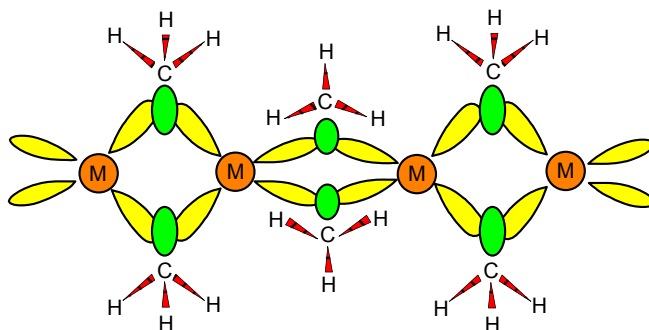


This compound and others of this type (referred to as metallocenes) will be discussed in greater detail in Chapter 22.

The structures of the organic derivatives of the Group IA and IIA metals are not simple due to the fact that many of them involve molecular association. For example, the lithium alkyls are tetramers in which the lithium atoms reside at the



**FIGURE 8.2** The structure of  $(\text{LiCH}_3)_4$ . There is a methyl group above each triangular face, but only two  $\text{CH}_3$  groups are shown for clarity. The overlap of the  $sp^3$  orbital on the methyl group with three  $2s$  orbitals on three lithium atoms is shown on the right.

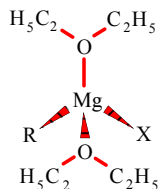


**FIGURE 8.3** The polymeric structure of  $\text{Mg}(\text{CH}_3)_2$  and  $\text{Be}(\text{CH}_3)_2$ . The two methyl groups in the middle of the drawing have the three hydrogen atoms directed above and below the plane of the page.

corners of a tetrahedron and the carbon atoms bonded to them are located above the triangular faces of the tetrahedron as shown in Figure 8.2.

In  $(\text{LiCH}_3)_4$ , the carbon atoms are above the centers of the triangular faces whereas in  $(\text{LiC}_2\text{H}_5)_4$  the bonded carbon is not above the center of the triangular face but rather it resides closer to one of the three lithium atoms on that face. The methyl compounds of beryllium and magnesium are polymeric with orbitals on the carbon atoms of the methyl groups forming three-center bonds or bridges that are similar to those formed in  $\text{B}_2\text{H}_6$  (see Chapter 9).

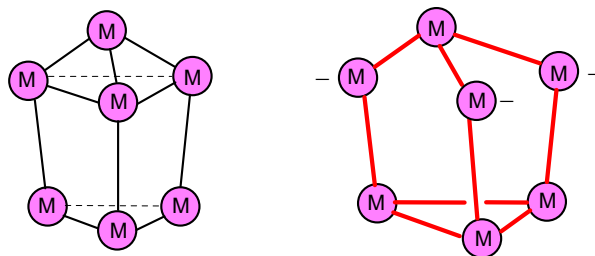
Grignard reagents are extensively associated, and  $\text{Mg}(\text{C}_2\text{H}_5)_2$  is polymeric with a structure similar to that shown in Figure 8.3. In solutions, molecular association occurs that is both solvent and concentration dependent. The degree of association also depends on the nature of the alkyl group and the halogen in the  $\text{RMgX}$  compounds. The compounds having the general formula  $\text{RMgX}$  also form solvates such as  $\text{RMgX} \cdot 2(\text{C}_2\text{H}_5)_2\text{O}$  for which the structure is



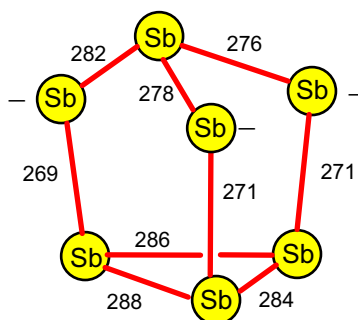
In this structure the bonding around the Mg is approximately tetrahedral. More details on the bonding in organometallic compounds of the Group IA and IIA metals will be presented in Chapter 22, but it must be stated again that these compounds are of great importance and utility in synthetic chemistry.

## 8.9 ZINTL COMPOUNDS

When lead is dissolved in a solution of sodium in liquid ammonia, a compound is formed between the two metals. Eduard Zintl (1898–1941) studied these compounds in Germany in the 1930s and is so generally considered as being the principle figure in such research that the compounds are now referred to as *Zintl compounds* or *Zintl phases*. Such compounds typically contain a metal from Group IA or IIA and element from Group IIIA, IV, or VA. It has been found that the anions in such compounds are  $\text{P}_7^{3-}$ ,  $\text{Pb}_9^{4-}$ ,  $\text{Sn}_7^{3-}$ , etc. Other compounds contain 14 atoms of the element found in later groups of the periodic table. For example,  $\text{Ba}_3\text{As}_{14}$ ,  $\text{Sr}_3\text{P}_{14}$ , and  $\text{Ca}_3\text{As}_{14}$  are such compounds.



**FIGURE 8.4** The structure of  $M_7^{3-}$  clusters (where  $M = P, As, Sb, \text{ or } Bi$ ). The structure on the left outlines the geometric structure whereas that on the right shows the bonds.



**FIGURE 8.5** The structure of the  $Sb_7^{3-}$  anion (adapted from Adolphson, Corbett, and Merryman, 1976) with bond lengths shown in pm.

Reacting a solution of the more active metal dissolved in liquid ammonia remains the general method of preparing Zintl compounds. However, some have been synthesized by direct combination of the elements. The structures of the anions that contain clusters are interesting. For example, the  $M_7^{3-}$  clusters have a trigonal prism structure with one triangular face capped as shown in Figure 8.4.

In terms of bonding, there are four atoms that have three bonds to them (and are thus considered to be neutral) whereas the other three atoms have only two bonds and are considered to be  $M^-$ . This gives the overall structure a charge of  $-3$ . In the compound  $[(2,2,2\text{-crypt})Na^+]_3Sb_7^{3-}$  (see Figure 8.1) the dimensions of the cluster anion (in pm) are as shown in Figure 8.5. Note that in the isolation of compounds containing the  $M_7^{3-}$  clusters the cation is large in accord with the hard–soft acid–base principle in which it is noted that large anions form stable lattices with large cations (see Chapter 6).

Although most of the members of the series have been known for some time,  $Bi_7^{3-}$  was a missing member of the  $M_7^{3-}$  series. Recently, Perla, Oliver, and Sevov (2014) isolated the compound  $[(2,2,2\text{-crypt})K^+]_3Bi_7^{3-}$ , which contains the  $Bi_7^{3-}$  ion. Thus, all of the  $M_7^{3-}$  cluster anions of P, As, Sb, and Bi have now been studied. The study of Zintl compounds continues to be an area of vigorous research in inorganic chemistry, and the active metals found in Groups IA and IIA are an integral part of that work.

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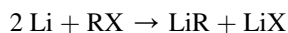
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## PROBLEMS

- Write complete, balanced equations to show the reactions of the following with water.
  - NaH
  - $\text{Li}_2\text{O}$
  - CaO
  - $\text{K}_2\text{S}$
  - $\text{Mg}_3\text{N}_2$
- Consider the reaction of sodium hydride in water. Describe what happens in terms of a Lewis acid–base reaction.
- Write complete, balanced equations to show the reactions of the following with water.
  - $\text{NaNH}_2$
  - $\text{CaC}_2$
  - $\text{Mg}_3\text{P}_2$
  - $\text{Na}_2\text{O}_2$
  - $\text{NaOC}_2\text{H}_5$
- Explain why it would be much more difficult to obtain a compound containing  $\text{Mg}^{2+}$  and  $\text{Mg}^{2-}$  than one containing  $\text{Na}^+$  and  $\text{Na}^-$ .
- Complete and balance the following.
  - $\text{C}_2\text{H}_5\text{OH} + \text{Na} \rightarrow$
  - $\text{Na}_2\text{O}_2 + \text{CO} \rightarrow$
  - $\text{NaHCO}_3 \xrightarrow{\text{heat}}$
  - $\text{KO}_2 + \text{H}_2\text{O} \rightarrow$
  - $\text{MgCl}_2 + \text{NaOH} \rightarrow$
- Use the data shown in Tables 4.1–4.3 to calculate the lattice energy for KBr. If the heats of hydration of  $\text{K}^+(\text{g})$  and  $\text{Br}^-(\text{g})$  are  $-322$  and  $-304 \text{ kJ mol}^{-1}$ , respectively, and  $\Delta H_{\text{vap}}$  of  $\text{Br}_2$  is  $30 \text{ kJ mol}^{-1}$ , calculate the heat of solution of  $\text{KBr}(\text{s})$  in water.
- Although the formation of  $\text{MgCl}$  from the elements would be accompanied by a release of energy,  $\text{MgCl}_2$  is the stable compound. Explain why  $\text{MgCl}$  is not a stable compound.
- Write complete equations for the following processes:
  - the preparation of calcium hydride
  - the use of calcium hydride as a drying agent
  - reaction between ethanol and butyl lithium
  - the preparation of lithium aluminum hydride
  - the preparation of beryllium metal by reduction (not by electrolysis)
- Complete and balance the following.
  - $\text{SrO} + \text{TiO}_2 \rightarrow$
  - $\text{Ba} + \text{H}_2\text{O} \rightarrow$
  - $\text{Ba} + \text{O}_2 \rightarrow$
  - $\text{CaF}_2 + \text{H}_2\text{SO}_4 \rightarrow$
  - $\text{CaCl}_2 + \text{Al} \rightarrow$
- Explain why  $\text{CP}_2^{2-}$  would be expected to be less stable than  $\text{CN}_2^{2-}$ .
- Describe the results of molecular weight studies on methyllithium and draw the basic structure of the molecule.
- Describe the structures and bonding in gaseous, liquid, and solid beryllium chloride.

13. Complete and balance the following.
- (a)  $\text{LiC}_2\text{H}_5 + \text{PBr}_3 \rightarrow$
  - (b)  $\text{CH}_3\text{MgBr} + \text{SiCl}_4 \rightarrow$
  - (c)  $\text{NaC}_6\text{H}_5 + \text{GeCl}_4 \rightarrow$
  - (d)  $\text{LiC}_4\text{H}_9 + \text{CH}_3\text{COCl} \rightarrow$
  - (e)  $\text{Mg}(\text{C}_5\text{H}_5)_2 + \text{MnCl}_2 \rightarrow$
14. Write complete equations for the following processes.
- (a) The preparation of butyl lithium
  - (b) The reaction of butyl lithium with water
  - (c) Dissolving of beryllium in sodium hydroxide
  - (d) Preparation of phenyl sodium
  - (e) Reaction between ethanol and lithium hydride
15. Molten beryllium chloride is not a good conductor of electricity, but when NaCl is dissolved in it, the solution becomes a good conductor. Write the appropriate equations and explain this observation.
16. In general, complexes of  $\text{Be}^{2+}$  are more stable than those of the heavier members of Group IIA. Explain this observation.
17. The reaction of lithium with an alkyl halide that takes place in an inert solvent can be shown as



Why would a reaction such as this be expected to take place readily?

## Chapter 9

# Boron

Boron is the forty-eighth most abundant element by weight in the earth's crust. This number is deceiving because boron ranks approximately thirtieth in the crust with respect to number of atoms. However, boron is not found in the elemental state in nature, but rather it is usually found as the tetraborate of sodium or calcium. The principal ore is *borax*,  $\text{Na}_2\text{B}_4\text{O}_7 \cdot 10 \text{H}_2\text{O}$ . Large deposits of borax are located in Southern California. It is here that about 75% of the world supply is mined.

The name borax is derived from the Persian word *borak*, meaning white. It has been used by man for centuries. Borax was used as a flux for soldering by the Egyptians, and this is still one of the more important uses for it. Borax was probably first imported to Europe by Marco Polo around 1300, and it was used in trade for many years afterward. However, borax was not used extensively until the deposits in California were discovered in the 1860s. Originally, the borax was removed from these deposits using mule teams, hence the name "20-mule team" borax.

Almost no other element has as much diversity in its chemistry as does boron. One of the main reasons for this is that boron shows a great tendency to form bonds to other boron atoms that results in complex cages and clusters. Because boron has three valence electrons, there is frequently formation of bonds that are more complex than the usual shared pair of electrons between two atoms. The chemistry is diverse also as a result of the existence of many binary borides and compounds in which a B–N group replaces a C–C unit. This is possible because each of these types of units has a total of eight valence shell electrons. This chapter will provide a survey of the chemistry of this interesting element.

### 9.1 ELEMENTAL BORON

Boron itself was first produced in an impure form by Sir Humphry Davy in 1808 by the electrolysis of molten boric acid. Gay-Lussac and Thenard also produced boron in 1808 by the reaction of potassium with boric acid. In 1895, Moissan produced boron by the reduction of  $\text{B}_2\text{O}_3$  with magnesium metal.



Boron prepared in this way contains magnesium and oxides of boron as impurities, and the purity ranges from 80 to 95% when the above procedure is used. The boron produced in this way is a crumbly brownish-black granular substance. This is known as the amorphous form of boron, and it has a density of  $2.37 \text{ g cm}^{-3}$ .

Small amounts of pure boron can be prepared by the reduction of boron trichloride by hydrogen on a heated tungsten filament.



Boron prepared in this way appears as black crystals having a density of  $2.34 \text{ g cm}^{-3}$ . The structure of the boron unit cell in the crystalline material is a regular icosahedral structure containing 20 equilateral triangles meeting at 12 vertices. Figure 9.1 shows the structure of the  $\text{B}_{12}$  icosahedron ( $I_h$  symmetry).

Boron cells may be considered as nearly spherical units and they can be arranged in more than one way in crystalline structures. The  $\text{B}_{12}$  units are surrounded by six others in a hexagonal arrangement with other layers similarly arranged lying above and below. These layers are stacked to give a three-dimensional structure. The crystalline forms of boron are tetragonal,  $\alpha$ -rhombohedral, and  $\beta$ -rhombohedral. The arrangement of  $\text{B}_{12}$  units in these structures is rather complex and they will not be shown here. However, all of the structures are extremely rigid, resulting in boron having a hardness of 9.3 compared to the value of 10.0 for diamond (Mohs' scale).

Naturally occurring boron consists of two isotopes,  $^{10}\text{B}$ , which comprises about 20% and  $^{11}\text{B}$ , which makes up the remaining 80%. This results in the average atomic mass being 10.8 amu.  $^{10}\text{B}$  has the ability to absorb slow neutrons to a great extent. Therefore, it finds application in nuclear reactors as control rods and protective shields. However, because

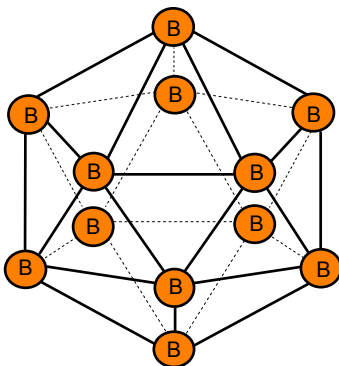


FIGURE 9.1 The icosahedral arrangement of the  $B_{12}$  structural unit.

boron itself is very brittle (and, therefore, nonmalleable) it must be combined or alloyed with a more workable material. Boron carbide is often mixed with aluminum and then processed into the desired shape.

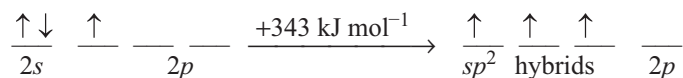
Research has been conducted on the use of the isotope  $^{10}\text{B}$  for treating brain tumors. It has been found that boron will tend to collect in the tumor to a much greater extent than in normal tissue. The tumor can then be exposed to bombardment with slow neutrons, and the  $^{10}\text{B}$  nuclei will emit alpha particles ( $^4\text{He}^{2+}$ ), which destroy the abnormal tissue.



Boron also has applications in metallurgy. Fibers of boron produced by depositing the element on tungsten wire can be coated with a resin to produce a strong, lightweight object (e.g., fishing rods, tennis rackets, etc.). By using this reinforced resin rather than aluminum as a construction material, there is a decrease of about 20% in the weight of the component, and its strength is comparable to that of steel.

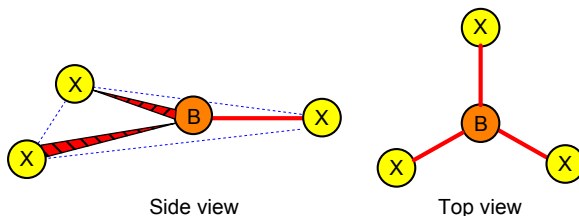
## 9.2 BONDING IN BORON COMPOUNDS

The electronic structure of the boron atom is  $1s^2 2s^2 2p^1$ . It might be expected that boron would lose the outer electrons and be present in compounds as  $\text{B}^{3+}$  ions. This ionization, however, requires over  $6700 \text{ kJ mol}^{-1}$ , and this amount of energy precludes compounds that are strictly ionic. Polar covalent bonds are much more likely, and it should be kept in mind that orbitals used in *bonding* are not necessarily the same as the orbitals in isolated atoms. *Hybridization* of atomic orbitals occurs and in the case of boron, the hybridization can be pictured as follows. Promoting a  $2s$  electron to one of the vacant  $2p$  orbitals can be accomplished with an increase in energy followed by hybridization of the orbitals to produce a set of  $sp^2$  hybrid *bonding* orbitals.



The energy necessary to promote the  $2s$  electron to a  $2p$  level is more than compensated for by the additional energy released when three equivalent bonds are formed.

From the above illustration, we expect boron to form three covalent bonds, which are equal in energy and directed  $120^\circ$  from each other. Accordingly, the boron trihalides,  $\text{BX}_3$ , have the following trigonal planar structure ( $D_{3h}$  symmetry).



In fact, all of the compounds containing boron bound to three other atoms have this configuration. In a few cases, such as  $\text{BH}_4^-$  and  $\text{BF}_4^-$ ,  $sp^3$  hybrids are formed and the species are tetrahedral ( $T_d$  symmetry).

### 9.3 BORON COMPOUNDS

When considering binary compounds of boron, it should be kept in mind that boron does not always behave as if the atom forms compounds in which the octet rule is obeyed. For example, compounds formed with scandium and titanium are  $\text{ScB}_2$  and  $\text{TiB}_2$ , but other compounds such as  $\text{Cr}_5\text{B}_3$  and  $\text{W}_2\text{B}_5$  also exist. Many of these borides have very complex structures. Several metal borides are known that have the formula  $\text{MB}_6$  where M is a metal having an oxidation state from +1 to +4. In most cases, the  $\text{B}_6$  units exist as octahedral clusters that occupy anion sites in a structure similar to that of cesium chloride. Other borides are known that contain  $\text{B}_{12}$  units in the crystal structures.

#### 9.3.1 Borides

Boron is rather unreactive but under certain conditions it forms one or more borides with most metals. For example, the reaction between magnesium and boron produces magnesium boride,  $\text{Mg}_3\text{B}_2$ .



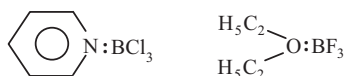
This product is hydrolyzed by acids to produce diborane,  $\text{B}_2\text{H}_6$ .



The fact that the expected product  $\text{BH}_3$  is not obtained will be discussed in a later section. Some metals form borides containing the hexaboride group,  $\text{B}_6^{2-}$ . An example of this type of compound is calcium hexaboride,  $\text{CaB}_6$ . In general, the structures of compounds of this type contain octahedral  $\text{B}_6^{2-}$  ions in a cubic lattice with metal ions. Most hexaborides are refractory materials having melting points over 2000 °C.

#### 9.3.2 Boron Halides

Boron halides, such as  $\text{BF}_3$  or  $\text{BCl}_3$ , are electron deficient molecules because they do not have an octet of electrons surrounding the boron atom. In accord with this property, they tend to act as strong Lewis acids by accepting electron pairs from bases to form stable *acid–base adducts*. Such electron donors as pyridine,  $\text{C}_5\text{H}_5\text{N}$ , or ether,  $(\text{C}_2\text{H}_5)_2\text{O}$ , will react with boron halides to give products such as



Because they are strong Lewis acids, the boron halides act as acid catalysts for several important organic reactions.

All of the boron halides are planar with bond angles of 120° in accord with our earlier description of  $sp^2$  hybridization. However, the B–X bond lengths found experimentally are shorter than the values calculated using the covalent single bond radii of boron and the halogens. This has been interpreted as indicating some double bond character due to  $\pi$ -bonding that occurs when electron density in filled  $p$  orbitals on the halogen atoms is donated to the empty  $p$  orbital on the boron (a situation known as back donation). The extent of this bond shortening can be estimated by means of the Shoemaker–Stevenson equation,

$$r_{\text{AB}} = r_{\text{A}} + r_{\text{B}} - 9.0 [\chi_{\text{A}} - \chi_{\text{AB}}] \quad (9.6)$$

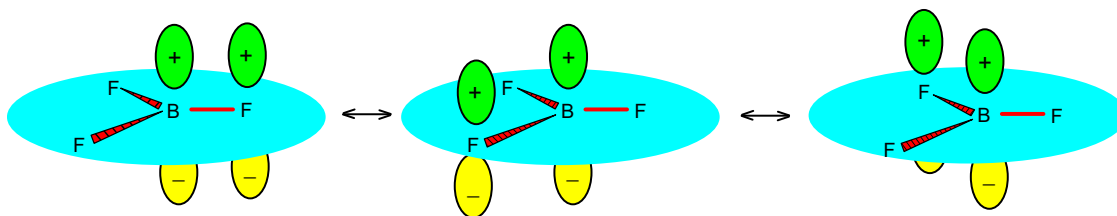
where  $r_{\text{AB}}$  is the bond length,  $r_{\text{A}}$  and  $r_{\text{B}}$  are the covalent single bond radii of atoms A and B, and  $\chi_{\text{A}}$  and  $\chi_{\text{B}}$  are the electronegativities of these atoms. The data shown in Table 9.1 show the application of this equation to the boron halides.

The results shown in Table 9.1 indicate that the extent of bond shortening is greatest for B–F bonds. This is to be expected because back donation should be more effective when the donor and acceptor atoms are of comparable size. Accepting of electron density in this way would be expected to reduce the tendency of the boron atom to accept electron density from a Lewis base. In accord with this, the strengths as acceptors toward the electron donor pyridine is

**TABLE 9.1** Bond Lengths in the Boron Halides

Bond type	Sum of covalent single bond radii (pm)	$r_{AB}$ Calculated from Eq. (9.6) (pm)	Experimental $r_{AB}$ (pm)
B–F	152	134	130
B–Cl	187	179	175
B–Br	202	195	187

$\text{BBr}_3 > \text{BCl}_3 > \text{BF}_3$ . The bond shortening discussed earlier is sometimes explained as being due to the contribution of resonance structures such as the following (adapted from House, 2013) where there is some degree of double bonding.



Boron halides also form some complexes of the type  $\text{BX}_4^-$ . For example,

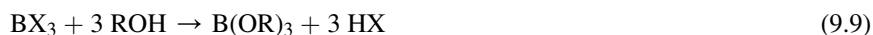


The  $\text{BCl}_4^-$  ion is tetrahedral because the addition of the fourth  $\text{Cl}^-$  means that an additional pair of electrons must be accommodated around the boron atom.

There are many reactions that boron halides undergo in addition to Lewis acid–base reactions. Typical of most compounds containing covalent bonds between a nonmetal and a halogen, these compounds react vigorously with water as a result of hydrolysis reactions. These reactions yield boric acid and the corresponding hydrogen halide, and they can be represented by the following general equation.



The  $\text{BX}_3$  compounds will also react with other protic solvents such as alcohols to yield borate esters.



Also, the reaction of  $\text{BCl}_3$  with  $\text{NH}_4\text{Cl}$  yields a compound known as trichloroborazine,  $\text{B}_3\text{N}_3\text{H}_3\text{Cl}_3$ , according to the equation



*Borazine*,  $\text{B}_3\text{N}_3\text{H}_6$ , will be described in more detail later in this chapter. Diboron tetrahalides,  $\text{B}_2\text{X}_4$ , are also known. These may be prepared in a variety of ways, among them the reaction of  $\text{BCl}_3$  with mercury,

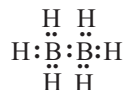


### 9.3.3 Boron Hydrides

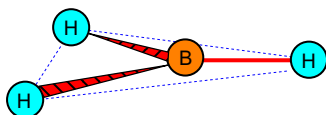
A large number of binary compounds containing boron and hydrogen are known. Generally referred to as the boron hydrides or hydroboranes, these compounds exist in an extensive array of structures. It might be assumed that the simplest of these compounds would be  $\text{BH}_3$ , borane, but instead a dimer,  $\text{B}_2\text{H}_6$  (diborane), is more stable. Several of these compounds were first prepared by Alfred Stock during the period of 1910–1930. They were prepared by the addition of hydrochloric acid to magnesium boride, an impurity found in boron produced during the reduction of  $\text{B}_2\text{O}_3$  with magnesium.



Stock found that the boron hydrides were volatile, highly reactive materials, and he devised special equipment and techniques for handling them. Six hydrides that he prepared are  $B_2H_6$ ,  $B_4H_{10}$ ,  $B_5H_9$ ,  $B_5H_{11}$ ,  $B_6H_{10}$ , and  $B_{10}H_{14}$ . Of particular interest to Stock was diborane, the structure of which was sometimes incorrectly shown as



However, in this structure 14 electrons are shown but the atoms have a total of only 12 valence electrons. It was also believed that borane,  $BH_3$ , should be a stable compound for which a structure for it could be drawn as



However, Stock found after numerous attempts that it was impossible to isolate borane. We now see that compound as being electron deficient, and that characteristic usually leads to some sort of molecular aggregation.

Even in a comparatively simple molecule such as  $B_2H_6$  elementary principles of bonding are inadequate to explain the structure. As a result of there being only 12 valence electrons, there cannot be the ordinary two-electron bonds throughout the molecule. Rather, the interpretation of the bonding in  $B_2H_6$  relies on the presence of two bonds known as three-center two-electron bonds. The structure of diborane is shown in Figure 9.2.

Another factor that is not accounted for in the incorrect valence bond structure is that the distance between the boron atoms is approximately that expected for a double bond. Four of the hydrogen atoms lie in the same plane as the boron atoms. The other two lie above and below this plane forming approximately a tetrahedral arrangement of four hydrogen atoms around each boron atom. The bridging hydrogen atoms do not form two covalent bonds, but rather simultaneously overlap with an  $sp^3$  hybrid orbital from each boron atom. This produces a so-called *three-center bond* (which can hold two electrons) for which the molecular orbital description is shown as in Figure 9.3.

In forming the wave function for the three-center bond, it is presumed that the wave function for the hydrogen atom is combined with a combination of boron wave functions. The bonds are considered to arise from the combination of two

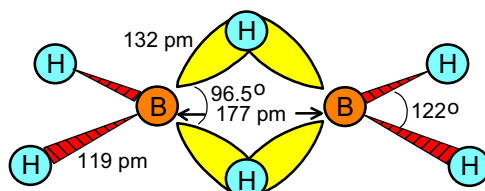


FIGURE 9.2 The structure of diborane in which bonds to boron atoms deviate somewhat from the expected tetrahedral angles.

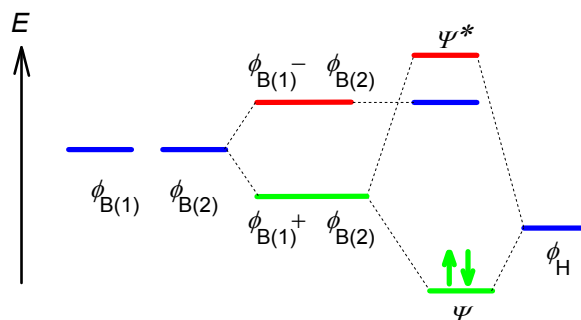
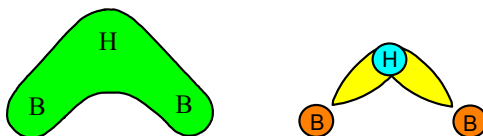


FIGURE 9.3 A molecular orbital diagram for a three-center B—H—B bond in diborane. In this diagram,  $\phi$  represents an atomic orbital, and  $\Psi$  and  $\Psi^*$  represent the bonding and antibonding three-center wave functions, respectively. Note the relative energies of the atomic orbitals on boron and hydrogen atoms as a result of the ionization potentials being 8.3 and 13.6 eV, respectively.

boron orbitals that are hybrids of  $s$  and  $p$  orbitals with a  $1s$  orbital on the hydrogen atom. Graphically, the formation of the three-center molecular orbital can be represented as shown below.



More complex boron hydrides, the polyhedral boranes, will be described in [Section 9.3.5](#).

Since the early work of Stock, other boron hydrides have been synthesized. Some of these compounds have been used as fuel additives, and they have found some application in high energy rocket fuels. However, as a result of  $B_2O_3(s)$  being one of the reaction products, the use of these materials in that way causes some problems. The boron hydrides will all burn readily to produce  $B_2O_3$  and water,



The properties of some boron hydrides along with those of other volatile hydrides are shown in Table 7.6.

### 9.3.4 Boron Nitrides

Boron has three valence shell electrons and nitrogen has five. Accordingly, the molecule BN is isoelectronic with  $C_2$ . Consequently, some of the allotropic forms that exist for carbon (graphite and diamond) also exist for materials having the formula  $(BN)_n$ . The form of  $(BN)_n$  having the graphite structure is very similar to graphite in many ways. Its structure consists of layers of hexagonal rings of boron and nitrogen atoms. Unlike graphite, the layers of boron nitride are not staggered. Rather, they fall directly in line with one another. The structures of graphite and boron nitride are shown in [Figure 9.4](#). Stronger van der Waals forces hold the sheets in line with each other so that boron nitride is not as good a lubricant as graphite. However, research is being conducted on the use of boron nitride as a high-temperature lubricant because of its chemical stability.

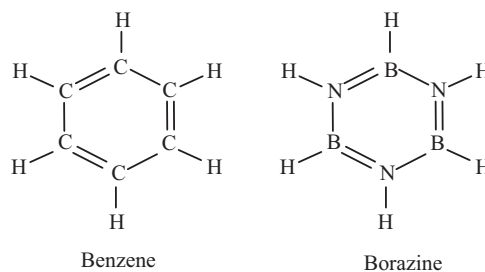
Boron nitride can be converted to a cubic form under conditions of high temperature and tremendous pressure. This cubic form is known as *borazon*, and it has a structure very similar to diamond. Its hardness is similar to that of diamond, and it is stable to higher temperatures. The extreme hardness results from the fact that the B–N bonds possess not only the covalent strength of C–C bonds, but also some ionic stabilization due to the difference in electronegativity between B and N. Borazon is not widely used at this time because of the difficulties in producing it.

Shortly after the discovery of carbon nanotubes in 1991, the synthesis of boron nitride nanotubes was reported. It was not until 2008, however, that monolayer and few-layer boron nitride analogues of graphene were isolated. The synthesis of single-layer boron nitride analogues of graphene can be carried out at high temperature by reaction of boric acid with urea to form boron nitride layers according to the equation



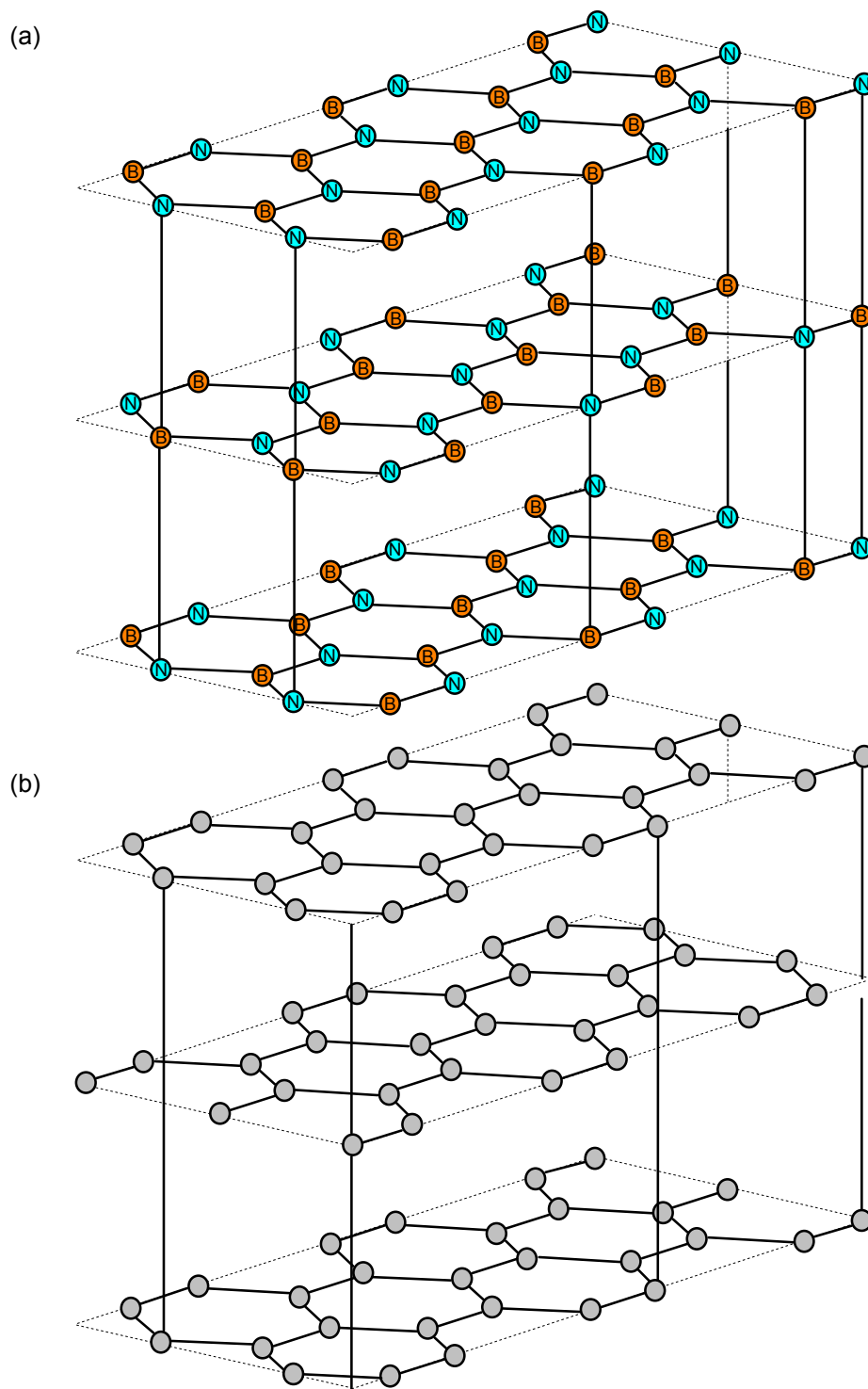
The number of layers that form can be controlled by varying the amount of urea.

Boron also forms many other compounds with nitrogen. One of the most interesting of these is *borazine*,  $B_3N_3H_6$  (m.p.  $-58^\circ C$ , b.p.  $54.5^\circ C$ ). The structure of borazine is very similar to that of benzene and, in fact, borazine has sometimes been referred to as “inorganic benzene.”



The borazine molecule has  $D_{3h}$  symmetry whereas benzene has  $D_{6h}$  symmetry. The length of the single B–N bond in  $H_3N-BF_3$  is 160 pm, but in borazine the B–N bond length is 144 pm. This is a reflection of the fact that in borazine there





**FIGURE 9.4** The structures of (a) boron nitride and (b) graphite (in which all atoms are carbon).

is substantial double bond character to the B–N bonds as a result of resonance. Although borazine resembles the aromatic benzene molecule in some respects, the electronic structure shows considerable difference. Theoretical studies show that although there is some delocalization in borazine, it is not as complete as in benzene. A description of this difference is beyond the scope of this book and the reader should consult the sources listed at the end of this chapter.

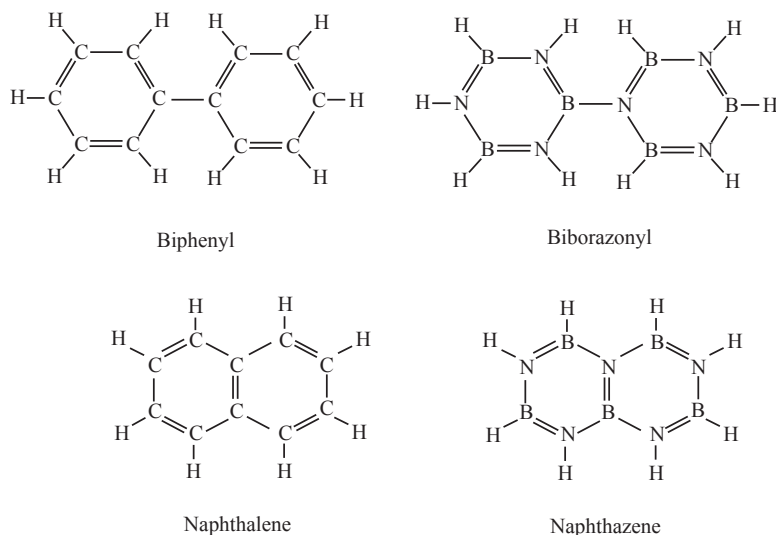


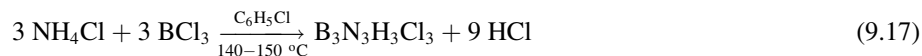
FIGURE 9.5 Structures of biphenyl and naphthalene and their BN analogues.

When heated under vacuum at high temperature, borazine loses hydrogen and polymerizes to yield products whose structures are similar to biphenyl and naphthalene. These structures are shown in Figure 9.5. Polymerization of this type involving the borazine ring is a process that may produce useful materials.

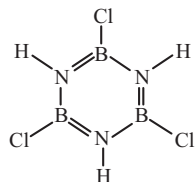
Borazine,  $B_3N_3H_6$ , was first prepared in 1926 as a product of the reaction of  $B_2H_6$  with  $NH_3$ .



It is usually prepared by the following reactions:



The compound  $B_3N_3H_3Cl_3$ , a trichloroborazine, has the structure

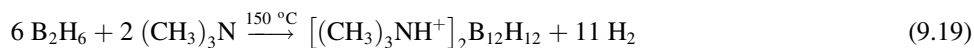


Note that in this structure the chlorine atoms are bonded to the boron atoms and the hydrogen atoms are attached to the nitrogen atoms. This arrangement rather than the opposite leads to stronger bonds as a result of the greater differences in electronegativity of the atoms involved in each type of bond. However, it should be kept in mind that boron halides will undergo many types of reactions that  $CCl_4$  will not. As a result, trichloroborazine is relatively reactive and the B—Cl bonds can be ruptured to give numerous derivatives that contain B—OH, B—OR, and other types of linkages. It should also be pointed out that although the  $\pi$ -electrons in benzene are distributed uniformly around the ring, this is not true in borazine compounds. As a result of the nitrogen atom having a greater attraction than boron for electrons, the electron density is more localized on the nitrogen atoms. This withdrawal of electrons from the boron atoms leaves them susceptible to attack by nucleophiles.

### 9.3.5 Polyhedral Boranes

A large number of boron compounds are known in which the boron atoms are arranged in the form of some sort of polyhedron (octahedron, square antiprism, bicapped square antiprism, icosahedron, etc.). In these structures, the boron

atoms are most often bonded to four, five, or six other atoms. The most common structure of this type is the icosahedron shown by the species  $B_{12}H_{12}^{2-}$ . This structure can be considered as a  $B_{12}$  icosahedron with a hydrogen atom bonded to each boron atom and the overall structure having a  $-2$  charge. It is prepared by the reaction of  $B_2H_6$  with a base that leads to the removal of hydrogen.



Compounds such as  $Cs_2B_{12}H_{12}$  are stable to several hundred degrees and do not behave as reducing agents as does the  $BH_4^-$  ion. An enormous number of derivatives of  $B_{12}H_{12}^{2-}$  have been prepared in which all or part of the hydrogen atoms are replaced by Cl, F, Br,  $NH_2$ , OH,  $CH_3$ ,  $OCH_3$ ,  $COOH$ , etc. These compounds and their reactions are far too numerous to be discussed individually here.

The structures of boranes can be grouped into several classifications. If the structure contains a complete polyhedron of boron atoms, it is referred to as a *closo* borane (*closo* comes from a Greek word meaning “closed”). If the structure has one boron atom missing from a corner of the polyhedron, the structure is referred to as a *nido* borane (*nido* comes from a Latin word for “nest”). In this type of structure, a polyhedron having  $n$  corners has  $(n - 1)$  corners that are occupied by boron atoms. A borane in which there are two corners unoccupied is referred to as an *arachno* structure (*arachno* comes from a Greek word for “web”). Other types of boranes have structures that are classified in different ways, but they are less numerous and will not be described.

The derivative of the  $B_{12}$  icosahedron that has the formula  $B_{12}H_{12}^{2-}$  has a hydrogen atom on each corner of a complete polyhedron so it represents a *closo* type of structure. Another example of this type is  $B_{10}H_{10}^{2-}$  that has eight boron atoms arranged in a square antiprism with an additional boron atom located above and below the top and bottom faces. Because it is possible for two boron hydrides having the same formula to have different structures, these terms are often incorporated in the name as prefixes. Thus, names such as *nido*- $B_5H_9$ , *arachno*- $B_4H_{10}$ , etc., are frequently used.

A relatively simple boron hydride that has an *arachno* structure is  $B_4H_{10}$  that has the structure shown in Figure 9.6(a). Pentaborane(9),  $B_5H_9$ , has a structure in which the boron atoms form a square-based pyramid with a terminal hydrogen atom attached to each boron atom and a bridging hydrogen atom along each edge of the square base. Because this structure, shown in Figure 9.6(b), can be considered as having an octahedron of boron atoms with one vertex vacant, it represents a *nido* type of borane.

The *closo*  $B_6H_6^{2-}$  structure is octahedral with a B–H group at each apex. The *closo*  $B_9H_9^{2-}$  ion has the structure shown in Figure 9.7(a), which is a trigonal prism having a capped structure above each rectangular face. The *closo* structure of  $B_{10}H_{10}^{2-}$  is a square antiprism having one B–H group capping each of the upper and lower planes as shown in Figure 9.7(b).

A derivative of  $B_{12}H_{12}^{2-}$  is the *carborane*,  $B_{10}C_2H_{12}$ . Note that this species is neutral because each of the carbon atoms has one more electron than does the boron atom and can, therefore, replace a B–H unit in the structure. Because all the positions in an icosahedron are identical, there will be three isomers of  $B_{10}C_2H_{12}$  that differ in the location of the two carbon atoms in the structure. The positions in the icosahedral structure are identified using the numbering system shown in Figure 9.8. Therefore, the structures for the three isomers of  $B_{10}C_2H_{12}$  are drawn as shown in Figure 9.9. The hydrogen atoms are not shown to simplify the drawing.

A derivative of the  $B_{10}C_2H_{12}$  species is the carborane anion  $B_9C_2H_{11}^{2-}$  that is missing a B–H group from one position in the icosahedron. This anion can function as a ligand in forming complexes with metals, and it forms a large number of metallocarboranes such as  $FeB_9C_2H_{11}$  in which  $Fe^{2+}$  is complexed with the cyclopentadiene anion,  $C_5H_5$ , to give a

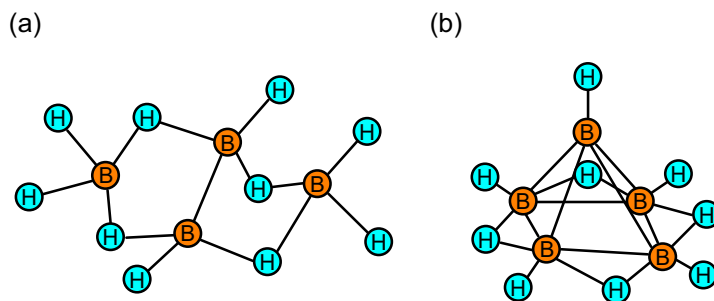
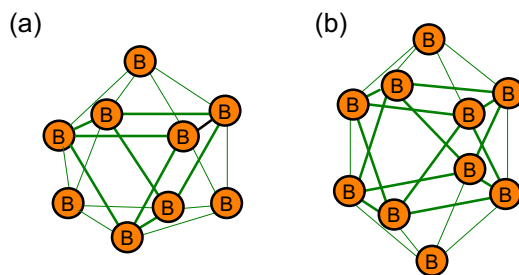
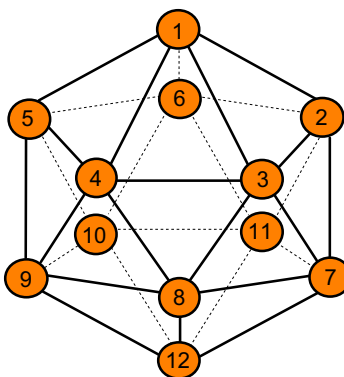


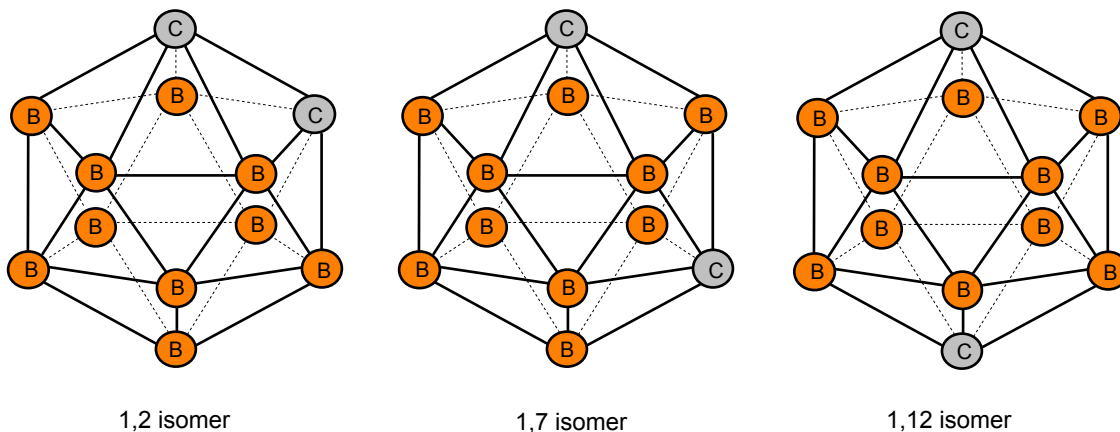
FIGURE 9.6 The *arachno* structure of  $B_4H_{10}$  (a) and the *nido* structure of  $B_5H_9$  (b).



**FIGURE 9.7** The *closo* structures of the  $B_9H_9^{2-}$  (a) and  $B_{10}H_{10}^{2-}$  (b) ions. In each case, there is one hydrogen atom attached to each boron atom, but the hydrogen atoms have been omitted to simplify the drawing.



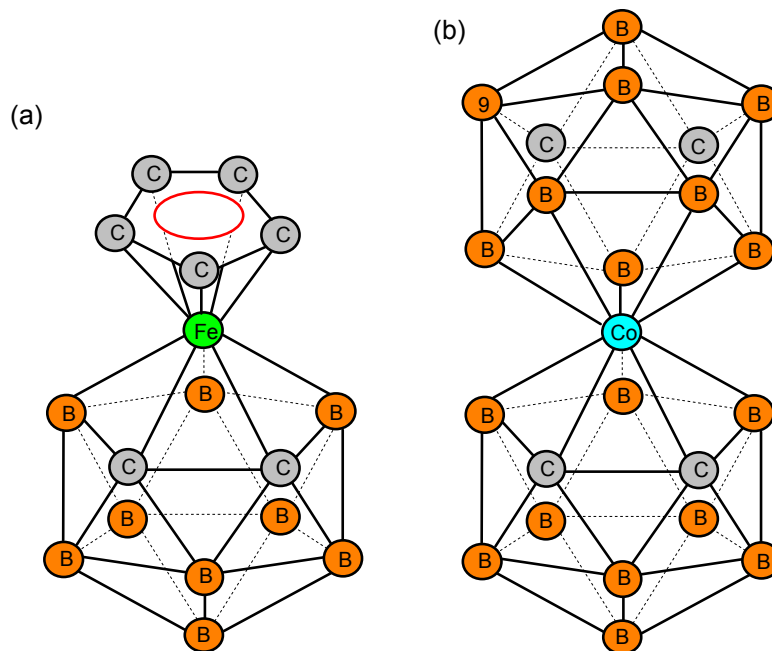
**FIGURE 9.8** The numbering system for an icosahedron such as that of the  $B_{12}$  molecule.



**FIGURE 9.9** The isomers of  $B_{10}C_2H_{12}$ . There is one hydrogen atom attached to each boron atom and each carbon atom. The hydrogen atoms have been omitted to simplify the drawing.

complex that has the structure shown in [Figure 9.10\(a\)](#). The complexes may contain one or two such ligands depending on the nature of the metal ion and the other ligands present. The complex with  $Co^{3+}$  has the formula  $[Co(B_9C_2H_{11})_2]^-$ , and its structure is shown in [Figure 9.10\(b\)](#).

An enormous number of derivatives of these basic polyhedral units (and also larger ones) exist and a great deal of their chemistry is known, but much of the chemistry of these compounds is outside the scope of this book and space limitations



**FIGURE 9.10** A complex of  $\text{Fe}^{2+}$  with cyclopentadiene and the  $\text{B}_9\text{C}_2\text{H}_{11}^{2-}$  ion (a) and the cobalt complex,  $[\text{Co}(\text{B}_9\text{C}_2\text{H}_{11})_2]^-$  (b). There is one hydrogen atom attached to each boron atom in these structures. The hydrogen atoms have been omitted to simplify the drawing.

prohibit a more complete discussion of them here. For more details, the references at the end of this chapter should be consulted.

Although boron forms a large number of unusual compounds, many of the better known ones are important and widely used. For example, the oxide  $\text{B}_2\text{O}_3$  is used extensively in the making of glass. Borosilicate glass, glass wool, and fiberglass are widely used because they are chemically unreactive, and can stand great changes in temperature without breaking. About 30–35% of the boron consumed is used in making types of glass. Borax has long been used for cleaning purposes in laundry products (detergents, water softeners, soaps, etc.). Boric acid,  $\text{H}_3\text{BO}_3$  (or more accurately,  $\text{B}(\text{OH})_3$ ), is a very weak acid that has been used as an eye wash. It is also used in flame retardants. Boron fiber composites are used in fabricating many items such as tennis rackets, aircraft parts, and bicycle frames. Borides of such metals as titanium, zirconium, and chromium are used in the fabrication of turbine blades and rocket nozzles. Although it is a somewhat scarce element, boron and some of its compounds are quite important and are used in large quantities.

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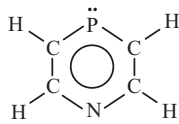
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## PROBLEMS

- Draw structures for the following molecules:
  - Borazine
  - The 1,2 isomer of  $B_{10}C_2H_{12}$
  - $B_2H_6$
  - Pentaborane(9)
- Write complete, balanced equations for the following processes:
  - The preparation of boron from the oxide
  - The preparation of  $B_2Cl_4$
  - The reaction of  $BCl_3$  with  $C_2H_5OH$
  - The preparation of borazine
  - The combustion of  $B_4H_{10}$
- Explain why  $BI_3$  is less stable than are the other boron halides.
- Explain why  $BF_3$  is a weaker Lewis acid than  $BCl_3$  toward a base such as  $NH_3$ .
- Write balanced equations for each of the following processes:
  - Preparation of  $B_3N_3H_6$
  - The reaction of  $B_3N_3Cl_3H_3$  with  $CH_3OH$
  - Preparation of  $(C_6H_5)_3B$
- On the basis of its structure, explain why boric acid is a weak acid that functions by complexing with  $OH^-$ . Draw the structure for the product.
- How would you predict that  $BF_3$  and  $B(CH_3)_3$  would interact with the following molecule?



- In the preparation of the adduct of  $B(CH_3)_3$  with ether, pyridine ( $C_5H_5N$ ) cannot be used as a solvent. In the preparation of the adduct of  $B(CH_3)_3$  with pyridine, ether can be used as the solvent. Explain the difference between the two cases.
- Would  $(C_2H_5)_3N$  or  $(C_2H_5)_3P$  form a stronger bond to  $BCl_3$ ? Explain your answer.
- In  $BF_3$ , the B–F bond length is 130 pm, but in  $BF_4^-$  it is approximately 145 pm. Explain this difference in the B–F bond lengths.
- Draw structures for the possible isomers of  $B_3N_3Cl_3H_3$ .
- Explain why in the  $B_3N_3H_3Cl_3$  molecule the chlorine atoms are bonded to boron atoms rather than to nitrogen.
- Would it be likely if  $BCl_3$  were dissolved in water that evaporating the water would allow  $BCl_3$  to be recovered? Write equations to explain your answer.
- Draw the structure for the boric acid molecule. It is possible for this compound to give a weakly acidic solution but without the loss of a proton. Show using an equation how this could occur.
- Suppose the molecule  $H_2PCH_2CH_2NH_2$  were to interact with one molecule of  $B(CH_3)_3$  and one of  $BCl_3$ . What would be the product? Explain your answer.

# Aluminum, Gallium, Indium, and Thallium

The elements below boron in Group IIIA of the periodic table include one of the most common and useful metals and three others that are much less important. Aluminum is the third most abundant element, and it occurs naturally in a wide variety of aluminosilicates, some of which will be described in more detail in Chapter 12. It also occurs in the minerals *bauxite*, which is largely  $\text{AlO}(\text{OH})$ , and *cryolite*,  $\text{Na}_3\text{AlF}_6$ . Although a few relatively rare minerals contain gallium, indium, and thallium, they are usually found in small quantities and are widely distributed. As a result, these elements are generally obtained as by-products in the smelting of other metals, especially zinc and lead.

## 10.1 THE ELEMENTS

Table 10.1 presents information on the sources of the metals in Group IIIA and Table 10.2 contains information on their properties and uses. All of these metals exhibit a +3 oxidation state in many of their compounds, but there is an increasing tendency toward a stable +1 state in heavier members of the group. As in the case of the chlorides of  $\text{Ag}^+$  and  $\text{Pb}^{2+}$ ,  $\text{TlCl}$  is insoluble in water and dilute  $\text{HCl}$ . In fact, the chemistry of thallium resembles that silver and lead in several respects.

There is also a pronounced tendency for the Group IIIA metals to form metal–metal bonds and bridged structures. The electron configuration  $ns^2 np^1$  suggests the possible loss of one electron from the valence shell to leave the  $ns^2$  pair intact. The electron pair that remains in the valence shell is sometimes referred to as an “inert” pair, and a stable oxidation state that is less than the group number by two units is known as an *inert pair effect*. The fact that oxidation states of +2, +3, +4, and +5 occur for the elements in Groups IVA, VA, VIA, VIIA, respectively, shows that the effect is quite common. Thus, it will be seen that the Group IIIA metals other than aluminum have a tendency to form +1 compounds, especially thallium.

Aluminum was discovered in 1825 by Hans Christian Oersted. Its name comes from the Latin word *alumen* for alum. Gallium was discovered by P. Boisbaudran in 1875, and its name is derived from the Latin name for France, *Gallia*. The existence of “eka-aluminum” (gallium) had been predicted earlier by Mendeleev, and its properties were estimated by comparison with those of its neighbors. Indium was found by F. Reich in zinc ores ( $\text{ZnS}$ ). Spectrographic analysis of the sulfide by T. Reich and H. Richter revealed lines having an indigo color from which the name for indium is derived. Thallium was discovered by Sir William Crookes who was studying slag from the production of sulfuric acid. His study was directed toward finding selenium and tellurium, but spectrographic analysis showed green lines characteristic of a new element, thallium. Its name is derived from the Greek word *thallos* meaning “green twig.”

**TABLE 10.1** Composition and Sources of the Group IIIA Minerals

Element	Minerals	Composition	Mineral Sources <sup>a</sup>
Al	Bauxite	$\text{AlO}(\text{OH})$ (Al, GA, AR)	Jamaica, Brazil, US
Ga	Sphalerite	$\text{ZnS}$	Canada, Mexico, Germany, US (MO, KS, OK, WI)
In	From Zn smelting	$\text{ZnS}$	Canada, Mexico, Germany, US (MO, KS, OK, WI)
Tl	From Pb and Zn smelting	$\text{PbS}$ , $\text{ZnS}$	Canada, Mexico, Germany, US (MO, KS, OK, WI)

<sup>a</sup>Not all of the minerals occur in each country.

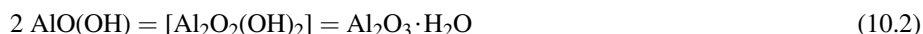
**TABLE 10.2** Properties and Uses of the Group IIIA Metals

	Production	Crystal Structure <sup>a</sup>	Density (g cm <sup>-3</sup> )	m.p., °C	b.p., °C	Uses
Al	Dissolve bauxite in NaOH, reform oxide, electrolysis	<i>fcc</i>	2.70	660	2467	Lightweight alloys, aircraft, food containers, foil
Ga	Recovered from smelting ZnS, electrolysis	<i>sc</i>	5.90	29.7	2403	Gallium arsenide LEDs, electronics
In	Recovered from smelting ZnS electrolysis	<i>tr</i>	7.31	157	2080	Electroplating, alloys
Tl	Recovered from smelting Pb and Zn	<i>hcp</i>	11.85	304	1457	

LED, light emitting diode.

<sup>a</sup>The structures indicated are as follows: *fcc* = face-centered cubic; *sc* = simple cubic; *hcp* = hexagonal close packing; *tr* = tetragonal.

Aluminum is obtained from bauxite by an electrolysis process that involves dissolving the mineral in molten *cryolite*, Na<sub>3</sub>AlF<sub>6</sub>. Bauxite is predominantly aluminum oxide, but there is convenient conversion between the oxide, hydrous oxide, and the hydroxide that can be represented as follows.



When Al(OH)<sub>3</sub> is precipitated from aqueous solutions, the product contains an indefinite amount of water, and it is correctly represented by a general formula Al(OH)<sub>3</sub>·*n*H<sub>2</sub>O or Al<sub>2</sub>O<sub>3</sub>·*x*H<sub>2</sub>O. Heating the product eventually results in the loss of all of the water to produce Al<sub>2</sub>O<sub>3</sub>. The hydrous oxide is not a true “hydrate” as is a compound such as CuSO<sub>4</sub>·5 H<sub>2</sub>O. The latter loses water in definite stages as can be revealed in a number of ways to produce CuSO<sub>4</sub>·3 H<sub>2</sub>O and CuSO<sub>4</sub>·H<sub>2</sub>O before becoming anhydrous CuSO<sub>4</sub>.

When the temperature of a sample of a hydrated material is increased at a constant rate and the mass of the sample is measured continuously, the behavior of true hydrates and hydrous oxides is different. A sample of a hydrous oxide tends to show a gradual decrease in mass as water is lost, but there are no well-defined stages representing definite compositions. In contrast, a true hydrate typically loses water in stages to give partially dehydrated materials that have definite compositions as in the case of CuSO<sub>4</sub>·5 H<sub>2</sub>O described above. In addition to following the loss of mass, other techniques can be used that show the formation of definite hydrates. For example, the vapor pressure of a compound such as CuSO<sub>4</sub>·5 H<sub>2</sub>O as a function of temperature also shows steps that correspond to the compositions CuSO<sub>4</sub>·5 H<sub>2</sub>O, CuSO<sub>4</sub>·3 H<sub>2</sub>O, and CuSO<sub>4</sub>·H<sub>2</sub>O. Monitoring the mass of a hydrous oxide such as Al<sub>2</sub>O<sub>3</sub>·*x*H<sub>2</sub>O as a function of temperature does not produce steps in the curve that correspond to fixed compositions. Therefore, although Al(OH)<sub>3</sub>, AlO(OH), and Al<sub>2</sub>O<sub>3</sub> exist, the role of water in the materials makes it likely that most naturally occurring solids containing them are mixtures. Bauxite actually represents a group of closely related minerals.

The electrolysis process for producing aluminum was developed by Charles Martin Hall while he was a student at Oberlin College. His work in aluminum chemistry eventually led to the formation of the Aluminum Company of America (ALCOA). The reduction process for aluminum can be represented as



This equation shows that 1 Faraday produces only one-third of a mole of aluminum metal. Because the atomic mass of aluminum is only approximately 27 g mol<sup>-1</sup>, the consumption of electricity is high for the production of a large quantity of the metal. Unfortunately, the selection of suitable chemical reducing agents for production of aluminum is quite limited, and recycling efforts are important.

In the production of aluminum, bauxite is treated with NaOH to produce NaAlO<sub>2</sub> (sodium aluminate) that reacts with HF to produce Na<sub>3</sub>AlF<sub>6</sub> (cryolite). The reaction can be represented as



In the electrolytic process, a mixture containing Al<sub>2</sub>O<sub>3</sub> dissolved in liquid Na<sub>3</sub>AlF<sub>6</sub> is electrolyzed.



No ores are sufficiently high in gallium, indium, or thallium content to be commercial sources of the metals. Accordingly, they are obtained as by-products of the smelting operations of zinc and lead. For example, gallium is obtained from *sphalerite* (ZnS) in the smelting of zinc. Indium is recovered from flue dust in zinc refineries, and thallium is isolated from by-products in lead and zinc refining. Gallium, indium, and thallium can be obtained by electrolysis of the solutions that are obtained by leaching the sulfides with acids. The oxides can also be reduced by carbon, hydrogen, or zinc.

Unlike the other Group IIIA metals in the +3 oxidation state,  $Tl^{3+}$  is a strong oxidizing agent, and many of its compounds react to produce  $Tl^+$  compounds when heated. The following equations illustrate this behavior.

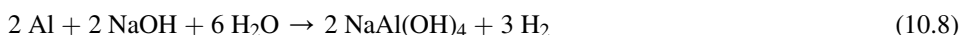
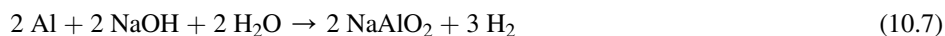


The  $Tl^{3+}$  ion is strong enough as an oxidizing agent to liberate oxygen from water.

There are no major uses of gallium, indium, and thallium as the metals themselves. Aluminum is a soft, silvery metal that is widely used as a structural metal. In most cases, it must be alloyed to give it the desired properties, and the pure metal has considerably less strength at temperatures above 300 °C than it does at low temperatures. The most useful and important alloys of aluminum are those that contain copper, magnesium, silicon, and manganese. An alloy that is used in cookware, gutters, etc., consists of aluminum and a small percentage of manganese.

Aluminum objects oxidize in air to form a thin oxide layer that is highly resistant to further oxidation and leaves the metal with a pleasing appearance. If it is so desired, aluminum objects can be given coatings in a variety of ways. When it is coated by an electrolytic process, it can be given a variety of colors. Composite materials of great strength can be obtained by fusing and bonding aluminum to boron or graphite fibers (see Chapter 11).

Aluminum is rapidly attacked by solutions of most acids, concentrated  $HNO_3$  being an exception. The fact that  $HNO_3$  is an oxidizing agent causes it to form an oxide layer on aluminum that makes it less reactive (it is said to become “passive”). However, aluminum also reacts violently with strong bases as shown by these equations.



Note that in these equations  $NaAl(OH)_4$  is formally equivalent to  $NaAlO_2 \cdot 2 H_2O$ . Aluminum will also displace hydrogen from water at high temperatures. Because of its ease of oxidation, aluminum powder is explosive in air.

As a metal, aluminum is a widely used and important metal. Approximately 70% of the aluminum consumed is used in building, transportation, packaging, and electrical conductors. It is used in construction of buildings, as an ornamental metal, making many types of machinery and tools, packaging material, catalysts, containers, and a wide variety of other applications. It is available as ingots, plates, rods, wire, and powder. Aluminum foil is extensively used for wrapping foods, and thin sheets are used as printing plates. Aluminum alloys are of enormous importance in the aircraft industry. It may be that the only reason the present age will not be referred to as the “Aluminum Age” is the importance of plastics.

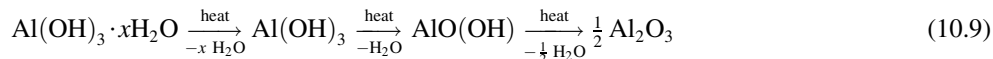
## 10.2 OXIDES

Of the oxides of the Group IIIA metals, by far the most important is aluminum oxide,  $Al_2O_3$ , because it is a widely used catalyst for many reactions. As described earlier in this chapter, there is an intimate relationship between the oxide and hydroxide in various degrees of hydration that result in the existence of the forms summarized below.

<i>gibbsite</i>	$\gamma$ -Al(OH) <sub>3</sub>
<i>bayerite</i>	$\alpha$ -Al(OH) <sub>3</sub>
<i>boehmite</i>	$\gamma$ -AlO(OH)
<i>diaspore</i>	$\alpha$ -AlO(OH) or $HAIO_2$

Bauxite consists of a group of closely related oxides and hydrated oxides, and it is a secondary mineral that results when silica is leached from minerals such as *kaolin*,  $Al_2Si_2O_5(OH)_4$ . The conditions for this type of leaching are favorable in the tropical areas in which bauxite is frequently found.

The relationship between the various forms of aluminum hydroxide, aluminum oxide, and the hydrous oxide is a complex one that depends on the method of preparation. From the standpoint of stoichiometry, the materials are related as shown in the following equation.



The behavior of aluminum oxide as a catalyst is strongly dependent on its treatment. The solid tends to take up water to form sites where an oxide ion is converted to a hydroxide ion by accepting a proton. There are also sites that are Lewis acids where exposed aluminum ions reside (electron deficient sites), and the oxide ions are Lewis bases. Because of these features, it is possible to some extent to prepare an aluminum oxide catalyst that has the necessary properties to match an intended use.

*Corundum*,  $\alpha\text{-Al}_2\text{O}_3$ , is a hard and chemically inert material that is a useful abrasive and refractory material that has the composition  $\text{Al}_2\text{O}_3$ . When traces of other metals are present, crystals of corundum take on colors that make them prized as gemstones. For example, *ruby* is corundum colored red by a small amount of chromium oxide. Many synthetic gemstones are produced by melting  $\text{Al}_2\text{O}_3$  to which is added a suitable metal oxide to impart the desired color.

Although  $\text{Al}_2\text{O}_3$  is the source of aluminum in the Hall process for obtaining the metal, it is a very useful material in its own right. Activated alumina is used as an adsorbent for many gases and is effective in their removal. Alumina is also an important constituent in abrasives and polishing compounds, catalysts, ceramics, and electrical insulators.

In addition to oxides based on the formula  $\text{Al}_2\text{O}_3$ , mixed oxides having the general formula  $\text{MM}'_2\text{O}_4$  are known where M is a doubly charged metal ion and M' is a triply charged metal ion. This combination of cations thus balances the total negative charge of  $-8$  produced by the four oxide ions. These compounds are known by the general name *spinel* after the mineral *spinel* that has the composition  $\text{MgAl}_2\text{O}_4$ . In terms of *composition*, this formula is equivalent to  $\text{MgO} \cdot \text{Al}_2\text{O}_3$ . Materials are known in which part or all of the  $\text{Mg}^{2+}$  ions can be replaced by other  $+2$  ions such as  $\text{Fe}^{2+}$ ,  $\text{Zn}^{2+}$ ,  $\text{Co}^{2+}$ ,  $\text{Ni}^{2+}$ ,  $\text{Be}^{2+}$ , or  $\text{Mn}^{2+}$ . For example, *gahnite* is  $\text{ZnAl}_2\text{O}_4$ , *hercynite* is  $\text{FeAl}_2\text{O}_4$ , and *galestite* is  $\text{MnAl}_2\text{O}_4$ . Madagascar and Sri Lanka are the usual sources of these spinels although some are found in Montana, New Jersey, and New York. Ghanite is also found in Massachusetts and North Carolina.

Numerous other oxides can have the same general formula as spinel. In fact, the formula  $\text{MM}'_2\text{O}_4$  can represent a large number of materials as a result of the total negative oxidation state of  $-8$  being balanced by several combinations of charges on M and M'. In the spinels discussed above, the charges are  $\text{M}^{2+}$  and  $\text{M}'^{3+}$  in which M represents a  $+2$  ion such as  $\text{Mg}^{2+}$ ,  $\text{Fe}^{2+}$ ,  $\text{Ca}^{2+}$ , etc. and M' represents a  $+3$  ion such as  $\text{Al}^{3+}$ ,  $\text{Fe}^{3+}$ ,  $\text{Cr}^{3+}$ ,  $\text{Ti}^{3+}$ , etc. However, if M is a metal having a  $+4$  oxidation state, M' can be a metal in the  $+2$  oxidation state to give compounds such as  $\text{PbFe}_2\text{O}_4$  and  $\text{TiMg}_2\text{O}_4$  in which Pb and Ti are  $+4$  and Mg and Fe are  $+2$ , respectively. In  $\text{MoNa}_2\text{O}_4$  the molybdenum is  $+6$  and Na is  $+1$ . The same general formula also represents compounds in which an anion having a  $-1$  charge is present. In that case, the total negative charge of  $-4$  can be balanced by one  $+2$  and two  $+1$  ions to give formulas such as  $\text{CaLi}_2\text{F}_4$ .

The spinel structure is based on a face-centered cubic arrangement of  $\text{O}^{2-}$  ions that results in cation sites that can be described as octahedral and tetrahedral holes. In the structure of spinel, the aluminum ions are surrounded by six oxide ions in an octahedral arrangement and magnesium ions are surrounded by four oxide ions in a tetrahedral arrangement. It should be noted that the  $+3$  ions reside in octahedral holes whereas the  $+2$  metal ions reside in tetrahedral holes (see Chapter 4). In a variation of this structure, compounds of this type form an arrangement in which half of the  $\text{M}'^{3+}$  ions reside in tetrahedral holes and the  $\text{M}^{2+}$  and half of the  $\text{M}'^{3+}$  ions reside in octahedral holes. Such a structure is known as an *inverse spinel*, and the general formula is sometimes written as  $\text{M} \cdot (\text{MM}')\text{O}_4$  to indicate the bonding sites of the ions. In a similar way, the *perovskite* structure for compounds having the general formula  $\text{M}^{\text{II}}\text{M}^{\text{IV}}\text{O}_3$  can also be considered as a case involving mixed metal oxides,  $\text{M}^{\text{II}}\text{O} \cdot \text{M}^{\text{IV}}\text{O}_2$  (see Chapter 4).

In Chapter 8, the effect of the charge to size ratio on the chemistry of ions was discussed, and the elements in Group IIIA provide some interesting observations. With the radius of  $\text{B}^{3+}$  being only 20 pm, the charge to size ratio is 0.15 which is quite high (for comparison, the values for  $\text{Be}^{2+}$  and  $\text{Al}^{3+}$  are 0.067 and 0.060, respectively). Such a high charge density polarizes anions and leads to bonding that is partially covalent. Therefore, compounds of B(III) and Be(II) are substantially covalent as are many compounds of Al(III). Boron oxide is slightly acidic whereas BeO and  $\text{Al}_2\text{O}_3$  are amphoteric. Gallium oxide is similar to aluminum oxide, but  $\text{Ge}_2\text{O}_3$  is slightly more acidic whereas  $\text{In}_2\text{O}_3$  and  $\text{Tl}_2\text{O}_3$  become more basic in that order. Thus,  $\text{Tl}_2\text{O}_3$  behaves only as a basic oxide.

In accord with the hard-soft interaction principle and the greater ease of reducing the metals, the oxides also become less stable for lower members of Group IIIA, and the order of thermal stability is  $\text{Al}_2\text{O}_3 > \text{Ga}_2\text{O}_3 > \text{In}_2\text{O}_3 \gg \text{Tl}_2\text{O}_3$ . It is also important to note that the heats of formation of the  $+3$  oxides are as follows:  $\text{Al}_2\text{O}_3$ ,  $-1670 \text{ kJ mol}^{-1}$ ;

$\text{Ga}_2\text{O}_3$ ,  $-1080 \text{ kJ mol}^{-1}$ ;  $\text{In}_2\text{O}_3$ ,  $-931 \text{ kJ mol}^{-1}$ ;  $\text{Tl}_2\text{O}_3$ ,  $-502 \text{ kJ mol}^{-1}$ . Although  $\text{Al}_2\text{O}_3$  can be heated to extremely high temperatures (m.p.  $\sim 2050 \text{ }^\circ\text{C}$ ),  $\text{Tl}_2\text{O}_3$  decomposes when heated to only about  $100 \text{ }^\circ\text{C}$ .



Being a much more nearly ionic compound,  $\text{Tl}_2\text{O}$  is a strong base (note that the formula resembles that of the Group IA metal oxides), and it reacts readily with water to give the hydroxide.



However,  $\text{Tl}_2\text{O}_3$  is not nearly as strongly a basic oxide, which is not surprising when it is recalled that covalent oxides tend to form acids. The reaction produced by heating thallium in air results in the formation of a mixture of  $\text{Tl}_2\text{O}$  and  $\text{Tl}_2\text{O}_3$ , but as mentioned earlier,  $\text{Tl}_2\text{O}_3$  is relatively unstable.

Aluminum oxide has such a high negative heat of formation that the reaction



is so exothermic that molten iron is obtained. This reaction, known as the *thermite reaction*, has been used to weld iron objects together by igniting a mixture of powdered aluminum and iron oxide at the junction of the metal objects. The high heat of formation of aluminum oxide also permits Al to be used as a reducing agent in the preparation of other metals, especially chromium where the reaction is



### 10.3 HYDRIDES

Although the hydrides having the general formula  $\text{MH}_3$  are known, only  $\text{AlH}_3$  is of much importance. Gallium hydride,  $\text{GaH}_3$ , is so unstable that it decomposes at room temperature. Aluminum hydride reacts violently with water



and it reacts explosively with air.



Lithium aluminum hydride,  $\text{LiAlH}_4$ , is a versatile and useful reducing agent that is widely used in organic chemistry. It is prepared by the reaction of excess lithium hydride with aluminum chloride in ether.



In a formal sense, this complex ion can be considered as a coordination complex in which  $\text{Al}^{3+}$  is surrounded by a tetrahedron consisting of four coordinated  $\text{H}^-$  ions that function as electron pair donors (see Chapter 6). In many of its reactions, the  $\text{AlH}_4^-$  behaves as if it contained  $\text{H}^-$  ions. For example, it reacts vigorously with water or alcohols as do ionic hydrides.



The extreme usefulness of  $\text{LiAlH}_4$  as a reducing agent in organic reactions is illustrated by the following general equations:



The reaction of  $\text{NaAlH}_4$  with  $\text{NaH}$  at elevated temperatures in an inert solvent produces  $\text{Na}_3\text{AlH}_6$ .



It is also possible to produce  $\text{Li}_3\text{AlH}_6$  directly from the metals and hydrogen at high pressure.



Although it is not actually an aluminum hydride, aluminum borohydride,  $\text{Al}(\text{BH}_4)_3$ , is a compound that is useful for reducing certain classes of organic compounds, and it has been used as an additive to fuels for jet engines. It can be prepared by the reaction



$\text{Al}(\text{BH}_4)_3$  undergoes spontaneous ignition in air and its reaction with water is violent.

## 10.4 HALIDES

Of the halogen compounds of the Group IIIA metals, the chlorides are by far the most important. Because the halogens are strong oxidizing agents, the usual product in the reaction of a halogen with a Group IIIA metal is the trihalide. Accordingly, the halides of Al, Ga, and In can be prepared by the reaction of the halogen with the metal. In the case of Tl, there is a strong tendency for the chloride compound to decompose to give  $\text{TlCl}$ . Because of the strong oxidizing nature of  $\text{Tl}^{3+}$ , it will oxidize iodide to  $\text{I}_2$ ,



and the iodine produced readily reacts with  $\text{I}^-$  to form  $\text{I}_3^-$ . Therefore, the compound having the formula  $\text{TlI}_3$  is actually  $\text{Tl}^+\text{I}_3^-$  rather than an iodide of  $\text{Tl}^{3+}$ . The reaction of chlorine with  $\text{TlCl}$  in hot aqueous slurries can be used to prepare  $\text{TlCl}_3$ , but solid  $\text{TlCl}_3$  decomposes at low temperature.

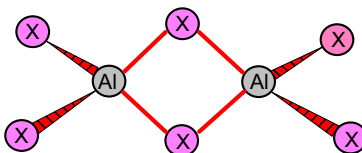


Thallium(III) reacts readily with halide ions to give complexes having the formula  $\text{TlX}_4^-$ . When  $\text{TlBr}_3$  is heated, it loses bromine to give a product having the formula  $\text{TlBr}_2$ .

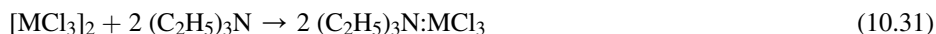
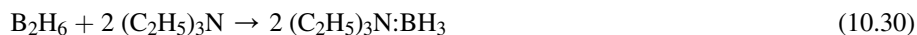


Studies indicate that  $\text{TlBr}_2$  (which appears to contain  $\text{Tl}^{2+}$ ) is actually  $\text{Tl}^+\text{TlBr}_4^-$ . Heating  $\text{GaCl}_3$  with metallic Ga produces “ $\text{GaCl}_2$ ”, a compound that has been shown to contain no unpaired electrons because it is diamagnetic. If the compound contained  $\text{Ga}^{2+}$ , it would be expected to be paramagnetic so this is another case of a compound containing a metal in two oxidation states,  $\text{Ga}^+\text{GaCl}_4^-$ . Compounds such as these illustrate the stability of the +1 oxidation state and the ability of the +3 ions to form stable complexes.

Because halides ions form bridges between two metal ions, the trihalides of aluminum exist as dimers in the vapor phase and in solvents that are poor electron donors. In general, the structures can be shown as



Although the bond angles are known to vary considerably depending on the metal and halogen, the arrangement is approximately tetrahedral around each metal atom. When the trihalides are dissolved in solvents that are Lewis bases, the dimers separate and complexes containing the monomer and the solvent are formed as a result of Lewis acid–base interactions. Such behavior is similar to that of borane and diborane in that the monomer,  $\text{BH}_3$ , is not stable but adducts of it are. This type of behavior is illustrated in the following equations.



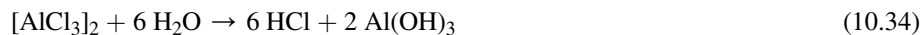
Of the halides of the Group IIIA metals,  $\text{AlCl}_3$  is the most important and widely used. It can be prepared by the reaction



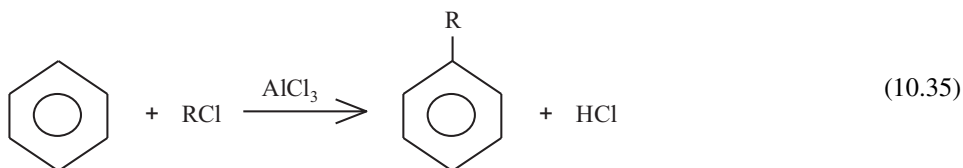
as well as by the reaction of the elements.



In many reactions,  $[\text{AlCl}_3]_2$  behaves as a covalent chloride, and it hydrolyzes readily in a vigorous reaction with water.



By far the most important use of aluminum chloride is as a catalyst in the *Friedel–Crafts reaction*. This use is derived from the fact that it is a strong Lewis acid and functions as an acid catalyst. Some of the most common reactions of this type involve the addition of alkyl groups to the benzene ring.

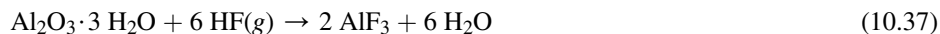


In this reaction,  $\text{AlCl}_3$  functions as a catalyst by generating a *positive* attacking species,  $\text{R}^+$  (see Chapter 6).



This alkylation reaction is of great importance in the preparation of ethyl benzene,  $\text{C}_6\text{H}_5\text{CH}_2\text{CH}_3$ , because it can be dehydrogenated to produce styrene,  $\text{C}_6\text{H}_5\text{CH}=\text{CH}_2$ . Styrene is polymerized to produce the enormous amounts of polystyrene and styrofoam that are fashioned into a large number of familiar objects.

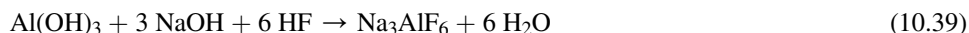
Although it is perhaps not used as extensively as aluminum chloride, aluminum bromide is also widely used as a Lewis acid catalyst. Aluminum fluoride is used in the preparation of *cryolite*,  $\text{Na}_3\text{AlF}_6$ , that is added to alumina to reduce its melting point and increase its electrical conductivity in the electrolytic production of aluminum. One reaction that can be employed to produce the fluoride is



The production of cryolite can also be carried out directly by the reaction of  $\text{AlF}_3$  and  $\text{NaF}$ .



In practice, an aqueous slurry of  $\text{Al}(\text{OH})_3$ ,  $\text{HF}$ , and  $\text{NaOH}$  is used and the reaction is represented by the equation



## 10.5 OTHER COMPOUNDS

The metals in Group IIIA react with several nonmetallic elements to produce interesting compounds. For example, aluminum nitride is produced by the reaction of the elements,



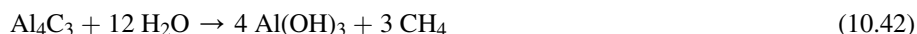
but it can also be produced by the reaction of  $\text{Al}_2\text{O}_3$  with carbon in a nitrogen atmosphere at high temperature as shown in the following equation.



Although they are not produced in this way,  $\text{GaN}$  and  $\text{InN}$  are known. Because the combination of a Group IIIA element with one from Group VA results in an average of four valence electrons for each atom, bonding similar to that between carbon atoms is suggested. Each atom should form four bonds that are oriented toward the corners of a tetrahedron. In accord with this prediction, the nitrides of Group IIIA elements have the wurtzite structure shown in Figure 4.5.

Compounds containing Group IIIA and Group VA elements include GaP and GaAs. Such compounds are isoelectronic with silicon, and like silicon, they behave as useful semiconductors. Gallium arsenide is used in the manufacture of light emitting diodes (LEDs). Aluminum arsenide is also used in electronic devices such as transistors, thermistors, and rectifiers. In general, compounds of the Group IIIA and VA elements are prepared by heating a mixture of the two elements.

In contrast to  $\text{CaC}_2$ , which is an acetylide that gives  $\text{C}_2\text{H}_2$  when it reacts with water, aluminum carbide,  $\text{Al}_4\text{C}_3$ , gives methane under these conditions.



Therefore, aluminum “carbide” is also considered to be aluminum methanide. It can be obtained by heating the elements or by the following reactions:



Aluminum carbide is used as a catalyst, a drying agent, and for generating small amounts of methane.

When aluminum dissolves in sodium hydroxide, the reaction can be represented by the equation



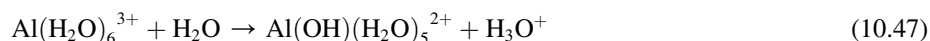
Evaporation of the resulting solution leads to the loss of water and produces a solid that has the composition  $\text{NaAlO}_2$ . The  $\text{AlO}_2^-$  ion is known as the *metaaluminate* ion and it can be regarded as the product obtained when water is lost from  $\text{Al}(\text{OH})_4^-$ . The reaction of Al with sodium hydroxide can also be written as



and the  $\text{AlO}_3^{3-}$  ion is known as the *orthoaluminate* ion, an ion that is analogous to silicates and phosphates (see Chapters 12 and 14). It can be viewed as resulting from the dehydration of the  $\text{Al}(\text{OH})_6^{3-}$  ion.

If a *metaaluminate* is precipitated as the salt containing a +2 cation such as  $\text{Mg}^{2+}$ , the formula can be written as  $\text{Mg}(\text{AlO}_2)_2$  or  $\text{MgAl}_2\text{O}_4$ , and this is the formula for the mineral *spinel*. The general spinel type of compound was discussed earlier, but these compounds, several of which were described, can be considered as the *metaaluminates* of +2 metals.

A class of aluminum compounds that has been known since ancient times is the *alums*. These compounds have the general formula  $\text{MAl}(\text{SO}_4)_2 \cdot 12 \text{H}_2\text{O}$  where M is a +1 cation. The more correct way to write the formula is  $[\text{M}(\text{H}_2\text{O})_6][\text{Al}(\text{H}_2\text{O})_6](\text{SO}_4)_2$ , which shows that the water molecules are attached to the metal ions. Because the aluminum ion hydrolyzes extensively in aqueous solutions, an acidic solution results when an alum is dissolved in water.

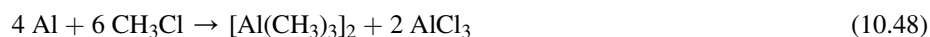


Aluminum hydroxide absorbs many dyes. Therefore, it is used as a mordant in dyeing processes because aluminum hydroxide attaches to cloth, especially cotton because the cellulose fibers have OH groups that can bind to  $\text{Al}(\text{OH})_3$ . An alum that is used as a mordant is  $\text{KAl}(\text{SO}_4)_2 \cdot 12 \text{H}_2\text{O}$ , but it is also employed in tanning leather, waterproofing compounds, and in many other applications. Aluminum sulfate,  $\text{Al}_2(\text{SO}_4)_3$ , is used in paper making, water purification, tanning leather, making fire extinguishing foams, and petroleum refining. Although they will not be dealt with here, many other aluminum compounds are of industrial importance.

## 10.6 ORGANOMETALLIC COMPOUNDS

Although there is a considerable organic chemistry of all of the Group IIIA metals, the organic chemistry of aluminum is by far the most extensive and important. A wide range of organic derivatives are known, and some of them are used industrially on a large scale. Several trialkylaluminum compounds are important as are some of the mixed alkyl halides. Table 10.3 shows physical data for some of the aluminum compounds.

Aluminum alkyls are prepared commercially by the reaction of aluminum with an alkyl halide as illustrated in the preparation of trimethylaluminum.



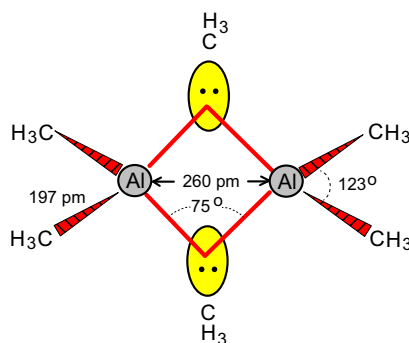
One of the interesting aspects of the chemistry of aluminum alkyls is the fact that they show a strong tendency to dimerize. In contrast,  $\text{GaR}_3$ ,  $\text{InR}_3$ , and  $\text{TlR}_3$  do not dimerize, but they are also less stable than their  $\text{AlR}_3$  analogs. The structure of the

**TABLE 10.3** Properties of Some Aluminum Alkyls

Compound <sup>a</sup>	b.p. (°C)	Density (g cm <sup>-3</sup> )	$\Delta H_{\text{vap}}$ (kJ mol <sup>-1</sup> )
Al(CH <sub>3</sub> ) <sub>3</sub>	126	0.752	44.9
Al(C <sub>2</sub> H <sub>5</sub> ) <sub>3</sub>	187	0.832	81.2
Al( <i>n</i> -C <sub>3</sub> H <sub>7</sub> ) <sub>3</sub>	193	0.824	58.9
Al( <i>i</i> -C <sub>4</sub> H <sub>9</sub> ) <sub>3</sub>	214	0.788	65.9
AlCl(C <sub>2</sub> H <sub>5</sub> ) <sub>2</sub>	208	0.958	53.7
AlCl <sub>2</sub> C <sub>2</sub> H <sub>5</sub>	194	1.232	51.9

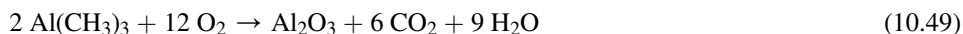
<sup>a</sup>Written as the monomer even though dimerization is extensive in some cases.

[Al(CH<sub>3</sub>)<sub>2</sub>]<sub>2</sub> dimer involves bridging as a result of overlap of aluminum orbitals with an *sp*<sup>3</sup> hybrid orbital on the methyl group to form a three-center bond. The structure of the dimer can be shown as



Trimethylaluminum and triethylaluminum are completely dimerized in the liquid phase. The alkyls containing longer chains or those having branched chains have a lower tendency to form dimers. There is evidence that the vaporization of [Al(C<sub>2</sub>H<sub>5</sub>)<sub>3</sub>]<sub>2</sub> leads to complete separation of the dimers but vaporization of [Al(CH<sub>3</sub>)<sub>3</sub>]<sub>2</sub> does not. The difference lies in the fact that the ethyl compound has a boiling point of 187 °C whereas that of the methyl compound is only 126 °C. Although the [Al(CH<sub>3</sub>)<sub>3</sub>]<sub>2</sub> dimers remain intact at 126 °C, they probably would not do so at 187 °C, the boiling point of the ethyl compound. The higher alkylaluminum compounds do not appear to dimerize completely even in the liquid state, and the dimers that do exist there appear to be completely dissociated in the vapor. The subject of association of aluminum alkyls will be treated in more detail in Chapter 22 as part of the overall discussion of metal alkyls.

Aluminum alkyls undergo a wide variety of reactions some of which involve the transfer of alkyl groups to other reactants. Only a brief introduction to this type of chemistry will be presented here because the chemistry of organometallic compounds will be discussed more fully in Chapter 22. Aluminum alkyls that contain methyl, ethyl, or propyl groups ignite spontaneously in air. For simplicity, the combustion reactions are written as being those of AlR<sub>3</sub> rather than [AlR<sub>3</sub>]<sub>2</sub>.



When small amounts of oxygen are used under controlled conditions, it is possible to convert the aluminum alkyl to the alkoxides.



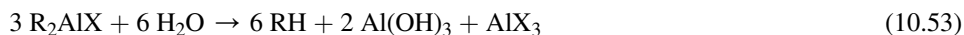
The alkoxides are easily hydrolyzed to produce alcohols.



The reactions of aluminum alkyls with water take place explosively when the alkyl group is C<sub>4</sub>H<sub>9</sub> or smaller. In reactions of this type, the alkyl behaves as if it contained R<sup>-</sup> that abstracts protons from polar bonds to hydrogen as shown in the following equation.



The mixed alkyl halides having the formula  $R_2AlX$  also react violently with water.



It is possible to carry out these reactions under less extreme conditions by preparing dilute solutions of the reactants in a nonhydroxylic solvent such as a hydrocarbon. The polar OH group of alcohols will also react with  $AlR_3$ .



As in the case of  $[AlCl_3]_2$ , the  $[AlR_3]_2$  dimers can be separated by reactions with molecules that behave as electron pair donors (nucleophiles). Typical nucleophiles that react this way are ethers and amines that give rise to the following reactions:



Many of the important reactions of the aluminum alkyls center on their ability to transfer alkyl groups to a variety of organic compounds. In many of these reactions, aluminum alkyls behave in much the same way as Grignard reagents ( $RMgX$ ) that are so important in organic chemistry. As illustrated by the following equations, they are also useful in producing other metal alkyls.



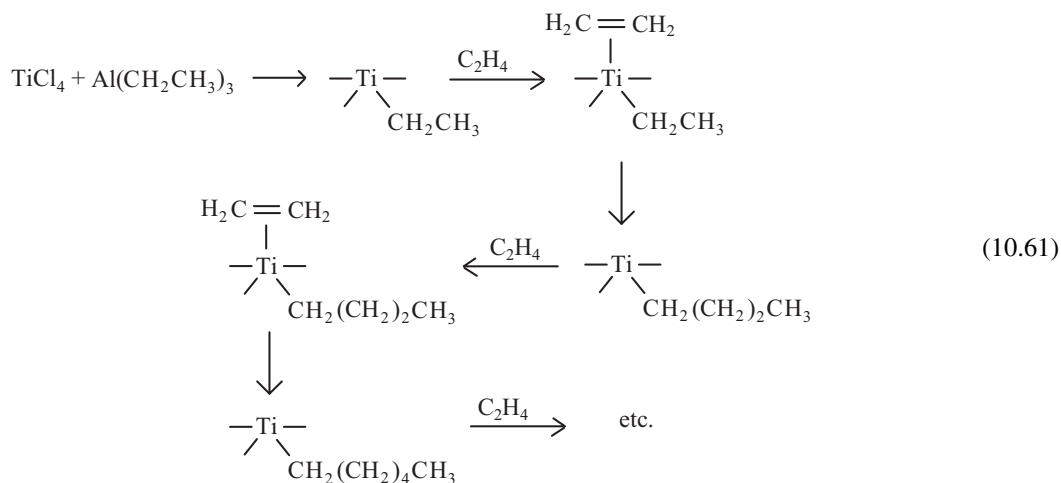
Aluminum alkyls are also employed in the synthesis of polymeric materials as shown in the following equation.



Another important use of aluminum alkyls is in the Ziegler–Natta polymerization process. This process for the polymerization of ethylene can be represented by the overall reaction



However, the mechanism is fairly complex, and it is usually illustrated as follows.



In this process,  $Al(C_2H_5)_3$  is employed to transfer ethyl groups to a titanium catalyst on a solid support. It is believed that ethylene then bonds to a vacant site on the titanium by functioning as an electron pair donor by means of the electrons in the  $\pi$  bond. The next step involves the migration of ethylene and its insertion in the  $Ti-C_2H_5$  bond to give a lengthened carbon chain. This is followed by the attachment of another  $C_2H_4$  to the vacant site on Ti and the process is repeated. Thus, the overall process represents the polymerization of ethylene using supported  $TiCl_4$  as a catalyst and triethylaluminum as the source of ethyl groups on the catalyst.



Only a brief survey of the vast area of organometallic chemistry of the Group IIIA metals has been presented, but it serves to show its broad applicability. Chapter 22 will deal with other aspects of the chemistry of some of the compounds described in this chapter.

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## PROBLEMS

1. An aqueous solution of aluminum sulfate is acidic. Write an equation to show how this acidity arises.
2. Provide an explanation of why  $\text{GaF}_3$  is predominantly ionic whereas  $\text{GaCl}_3$  is predominantly covalent.
3. Suppose a white solid could be either  $\text{Mg}(\text{OH})_2$  or  $\text{Al}(\text{OH})_3$ . What chemical procedure could distinguish between these solids?
4. How could  $\text{NaAlO}_2$  be prepared without the use of solution chemistry?
5. Why does the solubility of aluminum compounds that are extracted with sodium hydroxide decrease with the absorption of  $\text{CO}_2$  from the air?
6. Why is it logical that the  $+3$  ions would occupy the octahedral holes in the spinel structure?
7. Why is  $\text{AlCl}_3$  a dimer whereas  $\text{BCl}_3$  is monomeric?
8. When the molecular weights of  $\text{AlX}_3$  ( $X = \text{a halogen}$ ) are determined by freezing point depression in benzene solution, the molecular weights indicate that dimers are present. When the solvent used is ether, the molecular weights correspond to monomers of  $\text{AlX}_3$ . Explain this difference.
9. Write balanced equations for the following processes.
  - (a) The preparation of aluminum isopropylate (isopropoxide).
  - (b) The combustion of trimethyl aluminum.
  - (c) The reaction of  $\text{NaAlO}_2$  with aqueous  $\text{HCl}$ .
  - (d) The reaction of aluminum nitride with water.
10. The heat of formation of  $\text{GaCl}(g)$  is  $+37.7 \text{ kJ mol}^{-1}$  and that of  $\text{GaCl}_3(s)$  is  $-525 \text{ kJ mol}^{-1}$ . Show using equations why you would not expect  $\text{GaCl}(g)$  to be a stable compound.
11. Assume that an aluminum ore contains  $\text{Al}_2\text{O}_3$  and it is being leached out by aqueous  $\text{NaOH}$ . Write the equation for the process. Why would  $\text{Fe}_2\text{O}_3$  not dissolve under similar conditions?
12. Explain the trend shown in the text for the heats of formation of the  $\text{M}_2\text{O}_3$  compounds.
13. Prior to the development of the electrolytic process for the production of aluminum, sodium was used as the reducing agent for alumina. Write the equation for the process.
14. Why does  $\text{TlCl}_3$  contain  $\text{Tl}^{3+}$  whereas  $\text{TlI}_3$  does not?

## Chapter 11

# Carbon

Although carbon compounds form the basis of organic chemistry, there is a well-developed and important inorganic chemistry of carbon as well. In fact, the chemistry of C<sub>60</sub> and its derivatives, the fullerenes, is one of the most active new areas of inorganic chemistry. Prior to 1828, it was believed that all organic compounds came from living organisms. However, Friedrich Wöhler converted ammonium cyanate into urea,



and urea was considered to be an organic compound whereas ammonium cyanate was considered to be an inorganic compound. Thus, the classifications are now somewhat arbitrary.

Of the elements in Group IVA, carbon is a nonmetallic element that forms an acidic oxide, CO<sub>2</sub>,



Tin and lead are considered to be metals although silicon and germanium belong to the classification of metalloids. The elements in Group IVA show clearly the transition from distinct nonmetals to metal character encountered in progressing down in the group.

### 11.1 THE ELEMENT

Carbon and its compounds occur widely and the element is 14th in abundance. The most prevalent naturally occurring compounds are CO<sub>2</sub>, living organisms, coal, natural gas, and petroleum. Relatively small amounts of the element are found in the graphite and diamond forms.

Carbon occurs as <sup>12</sup>C (98.89%) and <sup>13</sup>C (1.11%). The interaction of neutrons produced by cosmic rays with <sup>14</sup>N in the atmosphere produces <sup>14</sup>C and protons,



The <sup>14</sup>C isotope has a half-life of 5570 years. As a result of the half-life being so long, a relatively constant amount of <sup>14</sup>C is contained in all carbon-containing materials. If a living organism dies, the <sup>14</sup>C is not replenished so it is possible to determine later the amount of <sup>14</sup>C present and thus determine the age of the material. Therefore, radiocarbon decay is frequently used as a dating procedure.

#### 11.1.1 Graphite

The outstanding characteristic of the chemistry of carbon is its ability to catenate. Catenation is the bonding of atoms of the same element to each other, and carbon shows this tendency to a far greater extent than any other element. Even its elemental forms are characterized by extended structural units. Graphite is the most common form of elemental carbon, but it exists in two forms both of which are composed of layers. One form has the layers oriented as shown in Figure 11.1. The structure shown in Figure 11.1 is the usual form of graphite, known as  $\alpha$ -graphite, that has carbon atoms aligned with others in the layers above and below to give the stacking pattern  $\cdots\text{ABAB}\cdots$ . The other structure is known as  $\beta$ -graphite in which the layers are staggered so that the carbon atoms in one layer lie above and below those in adjacent layers in a manner that gives a stacking pattern of  $\cdots\text{ABCABC}\cdots$ . It is possible to convert graphite from one form to the other.

Many of the properties of graphite result from its layered structure with mobile electrons. The bonding between layers is of the van der Waals type and is weak. Therefore, the layers slip over each other easily causing graphite to be a lubricant. The delocalized electron density within the fused ring system gives rise to the electrical conductivity of the material. Graphite is the standard state thermodynamic form of carbon so it is assigned a heat of formation of zero. The density of graphite is 2.22 g cm<sup>-3</sup>.

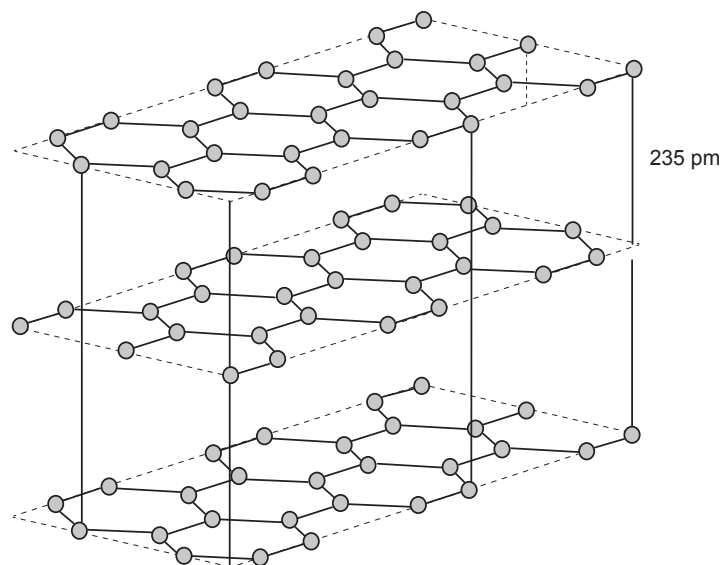


FIGURE 11.1 The layered structure of graphite.

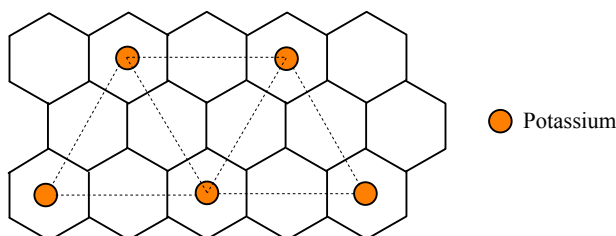


FIGURE 11.2 The view looking down on the layered structure of C<sub>8</sub>K.

Not only are the physical properties of graphite largely determined by its layered structure, but also its chemical reactivity is a result of that type of structure. Graphite reacts with potassium, rubidium, and cesium vapor at high temperature. The product obtained when potassium reacts with graphite is C<sub>8</sub>K in which layers of potassium atoms are located between layers of carbon atoms. Compounds of this type are referred to as *intercalation compounds*. The potassium atoms are arranged to give a hexagonal planar structure in which the atoms form triangles having corner atoms located below the center of a hexagon of carbon atoms. As shown in Figure 11.2, the spacing is such that there is a C<sub>6</sub> ring in each direction that has no potassium atom below its center.

The presence of the potassium atoms causes the distance between the layers of carbon atoms to increase from the value of 235 pm in graphite to 540 pm in C<sub>8</sub>K. When rubidium and cesium atoms are placed between the layers, the distances between the carbon layers are 561 and 595 pm, respectively. As would be expected for materials that contain atoms of an alkali metal, these materials are extremely reactive in air, and they react explosively with water. A large number of other intercalation compounds have been prepared that have halogens, interhalogens, or metal halides as the included substances.

### 11.1.2 Graphene

The individual two-dimensional sheets within graphite are referred to as *graphene*. Although the structure of graphite was elucidated in 1916, efforts to produce graphene, which is shown in Figure 11.3, were not successful until 2004 when Geim and Novoselov isolated layers that are single atom in thickness by micromechanical exfoliation. In graphene, the carbon atoms are arranged in hexagons with a carbon–carbon bond length of 142 pm. Graphene is semitransparent and flexible with a strength that is 100 times greater than that of steel. It is a zero-gap conductor of electricity and a good conductor of heat.

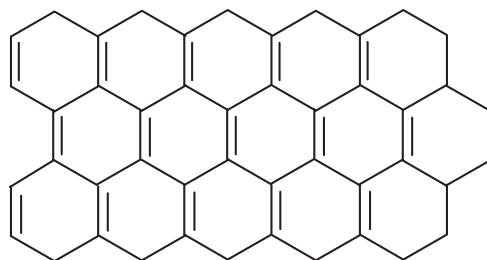


FIGURE 11.3 The structure of graphene.

Graphene can be synthesized by exposing layers of copper foil heated to 1800 °C to methane. The graphene is then obtained by etching away the copper with acid. Graphene can also be produced by exposing layers of platinum, nickel, or titanium carbide to benzene or ethylene at high temperatures. In 2010, Geim and Novoselov received the Nobel Prize in Physics for their work with graphene.

Single-walled carbon nanotubes are essentially sheets of graphene that are rolled into cylinders of approximately 1000 pm in diameter. Nanotubes with differing structures result depending on how the graphene sheet is twisted in the formation of the cylinder.

### 11.1.3 Diamond

Diamond has a cubic structure in which each carbon atom is bonded to four other carbon atoms, and it has no mobile electrons as in the case of the  $\pi$  electron system of graphite. The diamond structure is shown in Figure 11.4. Diamond, which has a density of  $3.51 \text{ g cm}^{-3}$ , is an insulator because the valence electrons are localized in single bonds. For the process



$\Delta H = 2.9 \text{ kJ mol}^{-1}$  at 300 K and 1 atm. Diamond is the hardest material known, although a compound containing boron and nitrogen, borazon, has almost equal hardness. Because of this property, diamond has a great many industrial uses in tools and abrasives. Because gem quality diamonds are prohibitively expensive for such uses, there has been a great deal of interest in producing diamonds synthetically. Industrially, this process is carried out under extreme conditions (3000 K and 125 kbar) to produce several tons of diamond annually. The process is catalyzed by molten metals such as Cr, Fe, or Pt, presumably due to the fact that graphite dissolves in the molten metal and diamond precipitates because of its lower solubility.

### 11.1.4 Fullerenes

The name *fullerene* comes from R. Buckminster Fuller, the designer of the geodesic dome. This form of carbon was obtained in 1985 by Smalley, Kroto, and coworkers by passing an electric arc between graphite rods in a helium atmosphere using a high current density. The process produces a soot, some of which is soluble in toluene. From this soluble soot, the principal product isolated is  $C_{60}$ , which has the structure shown Figure 11.5.

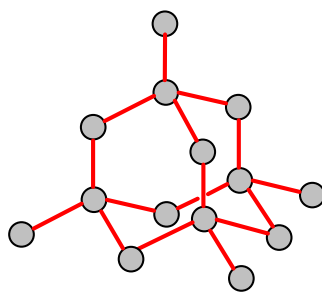


FIGURE 11.4 The structure of diamond.

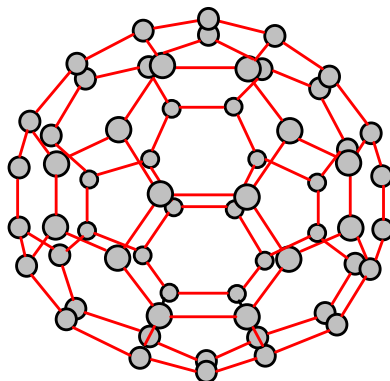


FIGURE 11.5 The structure of the  $C_{60}$  molecule.

Smaller amounts of  $C_{70}$  and other carbon aggregates are also produced. The presence of a  $C_{60}$  molecule was indicated by mass spectrometry as a result of a peak at 720 amu. Because the  $C_{60}$  structure has faces that are pentagons and hexagons joined along edges, it resembles the geodesic dome from the designer of which its name is derived.

In addition to  $C_{60}$ , a large number of other structures have been identified. These include  $C_{70}$ ,  $C_{76}$ ,  $C_{78}$ ,  $C_{84}$ , etc., but much less is known about these materials because they are much more difficult to produce and isolate. The chemistry of  $C_{60}$  has become extensive as a result of it being easier to produce and the fact that it is commercially available.

The fullerenes have shown a number of interesting properties and reactions. For example, it has recently been shown that  $C_{60}$  can be converted into diamond at room temperature when high pressure is applied. Interestingly,  $C_{60}$  is also a superconductor, and preparing it in the presence of metals can result in a metal atom being encapsulated in the  $C_{60}$  cage. Other materials have also been “shrink wrapped” inside the  $C_{60}$  molecule. Such complexes have been described as *endohedral*, meaning something is inside a cage or polyhedron. At this time, a number of complexes having catalytic properties have been prepared, such as  $((C_6H_5)_3P)_2PtC_{60}$  in which Pt bonds to two adjacent carbon atoms.

When  $C_{60}$  is treated with fluorine at elevated temperature, the reaction eventually leads to a product having the formula  $C_{60}F_{60}$  although fluorination under other conditions can lead to  $C_{60}F_{48}$ . It is possible to reduce  $C_{60}$  to produce negative ions known as *fullerides*. This can be accomplished electrochemically or by reactions with vapors of alkali metals. These compounds have the general formula  $M_nC_{60}$ , and some compounds of this type are superconductors. It has also been possible to attach organic groups to the carbon atoms.

With all of the multitude of reactions that have been carried out on  $C_{60}$  and related molecules, fullerene chemistry is growing at an incredible rate, and many interesting and potentially useful materials will undoubtedly be produced. For additional details on this type of chemistry, consult the references listed.

Other forms of carbon including charcoal, soot, lampblack, and coke are also known. Although these materials do not have structures that are highly regular, they are believed to have some local structure, and small units having graphite-like structures are known to exist. The fact that  $C_{60}$  was separated from soot shows that these useful and important forms of carbon are not completely without structure.

In addition to the production of diamonds described earlier, carbon is produced in two other important processes.



Furthermore, these “amorphous” forms of carbon can be transformed into graphite by means of the *Acheson process* in which an electric current is used to heat a rod of the “amorphous” form.

## 11.2 INDUSTRIAL USES OF CARBON

### 11.2.1 Advanced Composites

Composite materials are made up of two or more materials that have different properties. These materials are combined (in many cases involving chemical bonding) to produce a new material that has properties that are superior to either material alone. An example of this type of composite material is fiberglass in which glass fibers are held together by a polymeric resin.

The term *advanced composite* usually means a matrix of resin material that is reinforced by fibers having high strength such as carbon, boron, glass, or other material. The materials are usually layered to achieve the desired results. A common type of advanced composite is an epoxy resin that is reinforced with carbon fibers alone or in combination with glass fibers, in a multilayer pattern. Such composites are rigid, are light weight, and have excellent resistance to weakening (fatigue resistance). In fact, their fatigue resistance is better than that of steel or aluminum. Such properties make these materials suitable for aircraft and automobile parts, tennis racquets, golf club shafts, skis, bicycle parts, fishing rods, etc.

Fiber-reinforced composites may contain fibers of different lengths. In one type, the fibers are long, continuous, and parallel. In another type, the fibers may be chopped, discontinuous fibers that are arranged more or less randomly in the resin matrix. The properties of the composite are dependent upon its construction. Carbon fiber/resin composites have high strength to weight ratio and stiffness. Their chemical resistance is also high and they are unreactive toward bases. It is possible to prepare such materials that are stronger and stiffer than steel objects of the same thickness yet weigh 50–60% less.

Because carbon fibers of different diameters are available and because construction parameters can be varied, it is possible to engineer composites having desired characteristics. By varying the orientation, concentration, and type of fiber, materials can be developed for specific applications. The fibers can be layered at different angles to minimize directional characteristics. Also, layers of fibers can be impregnated with epoxy resin to form sheets that can be shaped prior to resin polymerization.

Because of their stiffness, strength, and light weight, carbon fiber composites are used in many applications in aircraft and aerospace fabrication, e.g., panels, cargo doors, etc. As a result of their high-temperature stability and lubrication properties, they are also used in bearings, pumps, etc. One drawback to the use of carbon fibers is the high cost of producing them. The cost can be as high as hundreds of dollars per pound. Military and aerospace applications have been the major uses of composite materials containing carbon fibers.

### 11.2.2 Manufactured Carbon

In addition to the use of carbon fibers in composites, there is extensive use of carbon in other manufacturing processes. For example, a mixture of coke and graphite powder can be prepared and then bonded with carbon. Typically, the carbon is added to the mixture in the form of a binder such as coal tar pitch or a resin. The mixture is put into the desired shape by compression molding or extrusion. Firing the object at high temperature (up to 1300 °C) in the absence of oxygen causes the binder to be converted to carbon, which holds the mass together. The wear resistance and lubricating properties of the finished part can be controlled by the characteristics and the proportions of coke, graphite, and binder.

The process described above produces an object that has pores. Metals, resins, fused salts, or glasses are frequently used in impregnation processes to fill the pores, a process that also alters the properties of the finished object. It is possible to machine these compact materials to close tolerances. The materials produced in this way are good conductors of heat and electricity. Because of the properties of the graphite present, they are also self-lubricating. Most solvents, acids, and bases do not attack the materials to a great extent. At high temperatures, oxygen slowly attacks these materials and they react slowly with oxidizing agents such as concentrated nitric acid. Although the carbon/graphite materials are brittle, they are stronger at high temperatures (2500–3000 °C) than they are at room temperature.

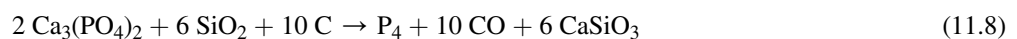
All of the properties described above make manufactured carbon a very useful material. It is used in applications such as bearings, valve seats, seals, dies, tools, molds, fixtures, etc. Specific uses of the final object may require materials that have been prepared to optimize certain properties. The fact that manufactured carbon can be prepared in the form of rods, rings, plates, tubes and other configurations makes it possible to machine parts of many types. Manufactured carbon represents a range of materials that have many important industrial uses.

### 11.2.3 Chemical Uses of Carbon

The chemical uses of elemental carbon are dominated by the fact that it is the cheapest reducing agent used on a large scale. Many processes illustrating this use are presented in this book. Two major ones are the production of iron,



and the production of phosphorus,



The reduction of metal compounds using carbon (charcoal) to produce metals has been known for many centuries. A problem associated with this use of carbon is that when an excess of carbon is used (which is necessary for a reaction to go to completion in a reasonable time), the product contains some carbon. It is not a “clean” reducing agent where the reducing agent itself and its oxidation product are gases. However, when cost is a factor, as it is in all large-scale industrial processes, carbon may be the reducing agent of choice. Another use of carbon depends on the fact that it adsorbs many materials. Particularly in the form of activated charcoal, carbon has a number of uses as an adsorbent.

## 11.3 CARBON COMPOUNDS

Carbon forms a wide variety of compounds with both metals and nonmetals. Compounds containing carbon in a negative oxidation state are properly called *carbides* and many such compounds are known. In a manner analogous to that of hydrogen (Chapter 7) and boron (Chapter 9), carbon forms three series of compounds, which are generally called *ionic*, *covalent*, and *interstitial carbides* although these types of bonding are not strictly followed. The carbides will be discussed under these headings even though there is a continuum of bond types. For example,  $\text{CaC}_2$  is essentially ionic but  $\text{SiC}$  is essentially covalent. As a result of metal oxides being reduced by carbon and forming important alloys containing carbon, numerous metal carbides are known that may represent stable phases in complex phase diagrams. For example, some of the species containing iron are  $\text{Fe}_2\text{C}$ ,  $\text{Fe}_3\text{C}$ , and  $\text{Fe}_7\text{C}_3$  and those of chromium are  $\text{Cr}_3\text{C}_2$ ,  $\text{Cr}_7\text{C}_3$ , and  $\text{Cr}_{23}\text{C}_6$ .

### 11.3.1 Ionic Carbides

If carbon is bonded to metals of low electronegativity, the bonds are considered to be ionic with carbon having the negative charge although this is an oversimplification. The metals include Groups IA and IIA, Al, Cu, Zn, Th, V, etc. Because carbon is in a negative oxidation state, these compounds react with water to give a hydrocarbon with water reacting as if it were  $\text{H}^+\text{OH}^-$ . However, some of these compounds give methane when they react with water (methanides) whereas others give acetylene (acetylides). Therefore, many ionic carbides often react as if they contain either  $\text{C}^{4-}$  or  $\text{C}_2^{2-}$  ions. In addition to these ions, the  $\text{C}_3^{4-}$  species is also known that has the linear structure common for 16-electron triatomic species.

The acetylide ion,  $\text{C}_2^{2-}$ , is isoelectronic with  $\text{N}_2$ ,  $\text{CO}$ , and  $\text{CN}^-$  (see Chapter 3), and it is presumed to be present in calcium carbide,  $\text{CaC}_2$ . Thus, a reaction producing acetylene from calcium carbide can be shown as



Because acetylene burns readily,



dripping water on  $\text{CaC}_2$  produces a supply of  $\text{C}_2\text{H}_2$  that can be burned as it is produced to serve as a portable light source. This is the basis of the “carbide light” or “miners lamp” that was in common use in the past by coal miners. Calcium carbide has a sodium chloride structure that is distorted in the vertical direction (a tetragonal structure) in which the  $\text{C}_2^{2-}$  ions are arranged between cations parallel to each other with a cation on either end as shown in Figure 11.6.

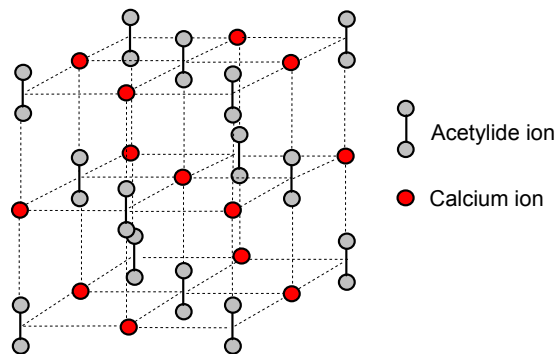
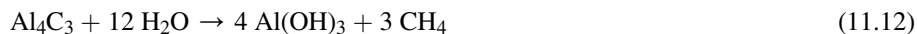


FIGURE 11.6 The structure of calcium carbide.

Acetylides can be produced in some cases by the direct reaction of carbon with a metal or metal oxide. For example,



Other carbides, e.g.,  $\text{Be}_2\text{C}$  and  $\text{Al}_4\text{C}_3$ , behave as if they contain  $\text{C}^{4-}$  and react with water to produce  $\text{CH}_4$ .



### 11.3.2 Covalent Carbides

When carbon forms compounds with other atoms having rather high electronegativity (Si, B, etc.) the bonds are considered to be covalent. The compounds formed, especially  $\text{SiC}$ , have the characteristics of being hard, unreactive refractory materials. Silicon carbide has a structure similar to diamond and it is widely used as an abrasive material. It is prepared by the reaction of  $\text{SiO}_2$  with carbon.



### 11.3.3 Interstitial Carbides

When most transition metals are heated with carbon, the lattice expands and carbon atoms occupy interstitial positions (see Chapters 7 and 9). The metals become harder, higher melting, and more brittle as a result. For example, when an iron object is heated and placed in a source of carbon atoms (charcoal or oils have been used historically), some iron carbide,  $\text{Fe}_3\text{C}$ , forms on the surface. If the object is cooled quickly (quenched), the carbide remains primarily on the surface and a durable layer results. This process is called *casehardening* and it was very important prior to the development of modern heat treatment processes for steels. Other carbides,  $\text{ZrC}$ ,  $\text{TiC}$ ,  $\text{MoC}$ , and  $\text{WC}$  are sometimes used in making tools for cutting, drilling, and grinding.

### 11.3.4 Oxides of Carbon

Carbon forms three well-known oxides. The simplest of these is carbon monoxide, which is isoelectronic with  $\text{N}_2$  and  $\text{CN}^-$ . It can be prepared by burning carbon in a deficiency of oxygen,



or by the reaction of C with  $\text{CO}_2$ ,



Because the oxides of nonmetals are acid anhydrides (see Chapter 6), CO is formally the anhydride of formic acid,



but this reaction cannot be carried out in this way. However, formic acid can be dehydrated to produce CO.



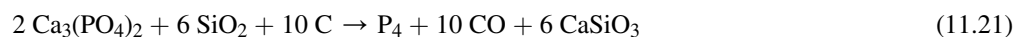
The reaction of CO with bases can be used to produce formates.



Carbon monoxide frequently results when carbon is used as a reducing agent. For example, the reactions



and





have been discussed earlier in this chapter. Another reaction that is used to produce large quantities of CO is

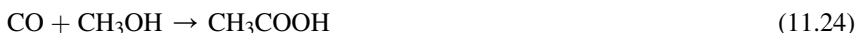


This is the basis for the water gas process and it was discussed in Chapter 7 as a method for preparing hydrogen.

Other important processes that lead to organic compounds utilize carbon monoxide. Two such processes are those used to produce methanol by the reaction

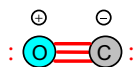


and acetic acid from methanol,



In addition to these uses, CO is employed in the production of  $\text{COCl}_2$ , phosgene, which is a widely used chlorinating agent that is usually produced on site. Phosgene has been used as a war gas so elaborate safety precautions are required.

The structure of CO can be shown as



and the carbon end of the molecule carries a negative formal charge. This end of the molecule has the “excess” of electron density that is held loosely so it is a soft electron pair donor. As a result, bonding is more favorable when the metal is uncharged or in a low oxidation state. The carbon end is the electron-rich end of the molecule and when carbon monoxide forms complexes with metals (metal carbonyls), it is the carbon end that binds to the metal.

A large number of metal complexes containing CO are known and most are formed by direct combination as illustrated by the following equations:



These compounds will be discussed further in Chapters 20 and 21. However, it should be mentioned here that many of the stable complexes have formulas that are based on the number of electrons needed by the metal to give the electron configuration of the next noble gas. For example, Ni has 28 electrons so four CO molecules each donating a pair of electrons gives Ni a total of 36 electrons, the number of electrons in Kr. Similarly, Fe ( $Z = 26$ ) and Cr ( $Z = 24$ ) need 10 and 12 electrons, respectively, so the stable complexes are  $\text{Fe}(\text{CO})_5$  and  $\text{Cr}(\text{CO})_6$ . In the case of Mn ( $Z = 25$ ), 11 electrons are needed. Five CO molecules contribute 10 electrons but there is still a deficiency of one electron and the manganese atom has one unpaired electron. Therefore, two such units combine to give  $(\text{CO})_5\text{Mn}-\text{Mn}(\text{CO})_5$  or  $\text{Mn}_2(\text{CO})_{10}$ .

Carbon monoxide is extremely toxic because it complexes with iron in the heme structure in the blood. When it binds to the iron, it binds more strongly than an  $\text{O}_2$  molecule so the CO destroys the oxygen carrying capability of the blood. Many deaths occur annually as a result of this behavior.

Carbon monoxide is a reducing agent and burns readily,



Perhaps the most important use of CO except for reduction of metals is in the production of methanol,



In this process, the catalyst consists of ZnO and Cu. Because methanol is an important solvent and fuel, this reaction is carried out on an enormous scale. In fact, methanol ranks about 22nd in the list of most used chemicals with an annual production of over 8 billion pounds.

The most familiar oxide of carbon is  $\text{CO}_2$ . It is most conveniently prepared by burning carbon in an excess of oxygen,

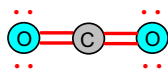


or by the reaction of a carbonate with a strong acid,



Solid  $\text{CO}_2$  sublimates at  $-78.5^\circ\text{C}$ . Because no liquid phase is present when this occurs, solid  $\text{CO}_2$  is called “dry” ice, and it is widely used in cooling operations.

The carbon dioxide molecule has a linear structure and is nonpolar (see Chapter 3).



It is the anhydride of carbonic acid,  $\text{H}_2\text{CO}_3$ , so solutions of  $\text{CO}_2$  are acidic.

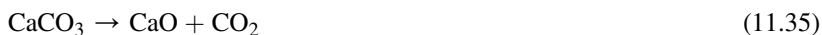


Also,  $\text{CO}_2$  will react with oxides to produce carbonates,



Many carbonates have important uses. One of the most important carbonates is calcium carbonate, a compound that is found in several mineral forms including *calcite*.

The most widely occurring form of calcium carbonate is limestone. This material is found throughout the world and it has been used as a building material for thousands of years. Heating most carbonates strongly results in the loss of carbon dioxide and the formation of an oxide. Heating  $\text{CaCO}_3$  strongly (called *calcining*) converts it into another useful material, lime. *Lime* ( $\text{CaO}$ ) is produced in this process, which is sometimes called “lime burning.”



Lime has been produced in this way for thousands of years. The loss of  $\text{CO}_2$  should result in a 44% mass loss, but in ancient times the material was considered ready for use when the mass loss was perhaps one-third of the original. Lime is used in huge quantities (about 38 billion lbs annually) in making mortar, glass, and other important materials, and to make  $\text{Ca}(\text{OH})_2$ , calcium hydroxide or *hydrated lime*. Even though  $\text{Ca}(\text{OH})_2$  is only slightly soluble in water, it is cheaper than  $\text{NaOH}$ , so it is widely used as a strong base.

*Mortar*, a mixture of lime, sand, and water, has been used in construction for thousands of years. The Appian Way, many early Roman and Greek buildings, and the Great Wall of China were constructed using mortar containing lime. In the Western Hemisphere, the early Incas and Mayans used lime in mortar for construction.

The actual composition of mortar can vary rather widely, but the usual composition is about one-fourth lime, three-fourths sand, and a small amount of water to make a paste. Essential ingredients are some form of solid such as sand and lime that is converted to  $\text{Ca}(\text{OH})_2$  by reaction with water.



Calcium hydroxide reacts with carbon dioxide from the atmosphere to produce calcium carbonate,  $\text{CaCO}_3$ ,



so that the grains of solid are held together by the  $\text{CaCO}_3$  to form a hard, durable mass. Essentially, artificial limestone is reformed.

*Concrete* is the man-made material used in the largest quantity. Concrete is made from inexpensive materials that are found throughout the world. In many cases, the raw materials are assembled at the construction site. Those raw materials are *aggregate* (sand, gravel, crushed rock, etc.) and *cement*, which bonds the aggregate together. Cement functions as the binding agent in the concrete. Particles of aggregate bond better when they have rough surfaces so that good adhesion occurs in all directions around each particle. A good mix should also have aggregate that is composed of a range of particle sizes rather than particles of a uniform size. A compact mass results with smaller particles filling the spaces between the larger ones. Basalt rock, crushed limestone, and quartzite are common materials used as aggregate. Lime is used in the manufacture of cement, the most common type being Portland cement, which also contains sand ( $\text{SiO}_2$ ) mixed with other oxides (aluminosilicates).

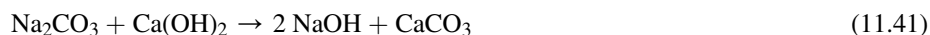
Portland cement is made by heating calcium carbonate, sand, aluminosilicates (see Chapter 12), and iron oxide to about  $870^\circ\text{C}$ . Kaolin, clay, and powdered shale are used as the sources of aluminosilicates. Heated strongly, this mixture loses water and carbon dioxide, and a solid mass is obtained. This solid material is pulverized and a small amount of calcium

sulfate ( $\text{CaSO}_4$ ) is added. Mortar used in modern times consists of sand, lime, water, and cement. Concrete usually consists of aggregate, sand, and cement. In both cases, cement binds the materials together by reacting with water to form a mixture known as *tobermorite* gel. This material consists of layers of crystalline material with water interspersed between them. To get strength, the correct amount of water must be used. Using too little water results in air being trapped in the mass giving a porous structure. If too much water is used, its evaporation and escape from the solid leaves it porous. In either case, the strength of the concrete suffers. Of course, the concrete can also be reinforced by using metal rods or wire. The complex reactions involved in the setting of concrete will not be discussed here, but it would be hard to overemphasize the importance of the use of calcium carbonate in concrete, mortar, and related materials.

Another important carbonate is sodium carbonate, which is also known as soda ash. The major source of soda ash today is from natural sources but prior to 1985, it was synthesized in large quantities. The synthetic process most often used is the *Solvay process*, and it consists of the following reactions:



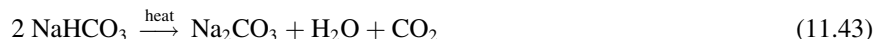
About 23 billion lbs of  $\text{Na}_2\text{CO}_3$  are produced annually. It is used in the manufacturing of glass (see Chapter 12), laundry products, water softeners, paper, baking soda, and sodium hydroxide by the reaction



In the United States, the major source of sodium carbonate is the mineral *trona*, which has the formula  $\text{Na}_2\text{CO}_3 \cdot \text{NaHCO}_3 \cdot 2 \text{H}_2\text{O}$ . The fact that sodium bicarbonate ( $\text{NaHCO}_3$ ) is present is no surprise in view of the fact that the carbonate reacts with water and carbon dioxide to produce sodium bicarbonate.

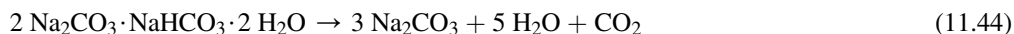


Heating the bicarbonate drives off water and  $\text{CO}_2$  to produce the carbonate,



The world's largest deposits of trona are found in Wyoming, but trona is also found in Mexico, Kenya, and Russia. The trona deposits in Wyoming are estimated to be 100 billion tons and they account for 90% of the US production. The trona mines in Wyoming alone account for 30% of the world production. In ancient times, sodium carbonate was obtained from the places where brine solutions had evaporated and dry lake beds. Such a naturally occurring material was not of high purity.

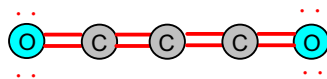
Trona is processed to obtain sodium carbonate by crushing it to produce small particles and then heating it in a rotary kiln. This dehydration process leaves impure sodium carbonate.



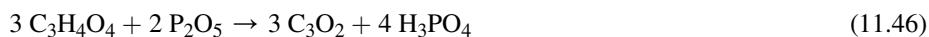
For many uses, the  $\text{Na}_2\text{CO}_3$  must be purified. This is done by dissolving it in water and separating shale and other insoluble material by filtration. Organic impurities are removed by adsorption using activated charcoal. The hydrated crystals of  $\text{Na}_2\text{CO}_3 \cdot \text{H}_2\text{O}$  are obtained by boiling off the excess water to concentrate the solution. The hydrated crystals are heated in a rotary kiln to obtain anhydrous sodium carbonate.



Carbon suboxide,  $\text{C}_3\text{O}_2$ , contains carbon in the formal oxidation state of  $+4/3$ . Because this is lower than its oxidation state in  $\text{CO}$  or  $\text{CO}_2$ , the oxide is called carbon suboxide. It could also be named as tricarbon dioxide. The linear structure of the molecule can be shown as



The molecule contains three carbon atoms bonded together, so this suggests a method for its preparation by dehydrating an organic acid containing three carbon atoms. Because  $C_3O_2$  is formally the anhydride of malonic acid,  $C_3H_4O_4$  (also written as  $HOOC-CH_2-COOH$ ), one way of preparing  $C_3O_2$  is by the dehydration of the acid.



As expected, the reaction of  $C_3O_2$  with water produces malonic acid. Carbon suboxide burns readily and although it is stable at  $-78^\circ C$ , it polymerizes at  $25^\circ C$ .

### 11.3.5 Carbon Halides

Of the halogen compounds of carbon, the most important is  $CCl_4$ , which is widely used as a solvent. However, the fully halogenated compounds are usually considered as derivatives of methane, and they are usually considered as organic in origin. Consequently, they will not be described in much detail here.

The reaction



can be used to prepare  $CCl_4$  (b.p.  $77^\circ C$ ). The  $S_2Cl_2$  obtained has many uses, including the vulcanization of rubber (see Chapter 16). The reaction of methane with chlorine also produces  $CCl_4$ .



Although it is widely used as a solvent,  $CCl_4$  is known to constitute a health hazard and its widespread use in dry cleaning has been largely halted.  $CCl_4$  does not hydrolyze in water as do most covalent halogen compounds.

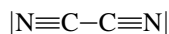
### 11.3.6 Carbon Nitrides

The only significant compound of carbon and nitrogen is cyanogen,  $(CN)_2$ . The cyanide ion,  $CN^-$ , is a *pseudohalide* ion, which means that it resembles a halide ion because it forms an insoluble silver compound and it can be oxidized to the  $X_2$  species. Cyanogen was first obtained by Gay-Lussac in 1815 by heating heavy metal cyanides.



It can also be prepared from carbon and nitrogen by electric discharge between carbon electrodes in a nitrogen atmosphere.

The structure of  $(CN)_2$  is



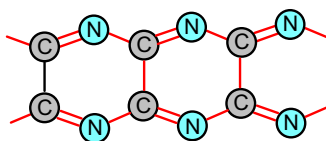
with  $N \equiv C$  bond lengths of 116 pm and a  $C-C$  bond length of 137 pm. It is a colorless gas that is highly toxic. Combustion produces a violet flame with the products being  $CO_2$  and  $N_2$ .



When heated, cyanogen polymerizes to give a white solid known as *paracyanogen*.



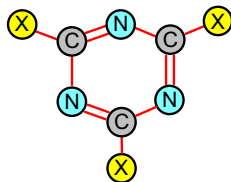
This material has the structure that is shown below.



A large number of other compounds containing carbon and nitrogen are known. For example,  $NC-CN$  is cyanogen, but  $CN-CN$  is known as isocyanogen. Although they are considered to be organic compounds, cyanides or nitriles,  $RCN$ , and isocyanides or isonitriles,  $RNC$ , are known. Because  $CN^-$  is a pseudohalide, compounds that contain this group bonded to a halogen are equivalent to an interhalogen (actually, a pseudo interhalogen). These compounds undergo the reaction



giving products having the following structure.



A similar reaction occurs when  $\text{X} = \text{NH}_2$  to yield  $(\text{H}_2\text{NCN})_3$ , a compound known as melamine.

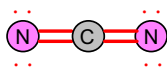
Although the chemistry of compounds containing only carbon and nitrogen is quite limited, many cyanides have widespread use. Calcium cyanamide,  $\text{CaCN}_2$ , can be prepared by the reaction



This process is important because it represents a simple way for nitrogen fixation (forming compounds from atmospheric nitrogen). The  $\text{CaCN}_2$  is used as a fertilizer because the reaction



produces both  $\text{NH}_3$  and  $\text{CaCO}_3$ . This use is not as extensive as it was in former times (see Chapter 14). The structure of the  $\text{CN}_2^{2-}$  ion, which like  $\text{CO}_2$ ,  $\text{OCN}^-$ , and  $\text{SCN}^-$  contains 16 electrons, can be shown as



Cyanamides can be converted to cyanides by reaction with carbon,



Cyanides are extremely toxic and acidifying a solution containing  $\text{CN}^-$  produces  $\text{HCN}$ .



Hydrogen cyanide (b.p.  $26^\circ\text{C}$ ) is a very toxic gas. It is a weak acid ( $K_a = 7.2 \times 10^{-10}$ ) so solutions of ionic cyanides are basic due to hydrolysis.



The  $\text{CN}^-$  ion is a good coordinating group, and it forms many stable complexes with metals (see Chapters 20 and 21).

Cyanates can be prepared from cyanides by oxidation reactions. For example,



Cyanides will react with sulfur to produce thiocyanates.



$\text{HSCN}$  is a strong acid (comparable to  $\text{HCl}$ ) and it can readily form amine hydrothiocyanates,  $\text{R}_3\text{NH}^+\text{SCN}^-$ , which are analogous to amine hydrochlorides. In a matter similar to the behavior of  $\text{CN}^-$ ,  $\text{SCN}^-$  forms many complexes with metal ions, and it bonds to soft metals (e.g.,  $\text{Pt}^{2+}$  or  $\text{Ag}^+$ ) through the sulfur atom and to hard metals (e.g.,  $\text{Cr}^{3+}$  or  $\text{Co}^{3+}$ ) through the nitrogen atom (see Chapter 6).

### 11.3.7 Carbon Sulfides

The most common compound of carbon and sulfur is  $\text{CS}_2$ , carbon disulfide. It can be prepared by the reaction of carbon and sulfur in an electric furnace or by passing sulfur vapor over hot carbon.



$\text{CS}_2$  (b.p.  $46.3^\circ\text{C}$ ) is a good solvent for many substances including sulfur, phosphorus, and iodine. The compound has a high density ( $1.3 \text{ g ml}^{-1}$ ) and it is slightly soluble in water, although it is completely miscible with alcohol, ether, and benzene. It is also quite toxic and highly flammable. It is used in the preparation of  $\text{CCl}_4$  as described earlier in this chapter. One interesting reaction of  $\text{CS}_2$  is analogous to the behavior of  $\text{CO}_2$ .



The  $\text{CS}_3^{2-}$  ion, which like  $\text{CO}_3^{2-}$  has a  $D_{3h}$  structure, is known as the thiocarbonate ion. In addition to thiocarbonates, other anionic species include  $\text{C}_3\text{S}_3^{2-}$ ,  $\text{C}_4\text{S}_4^{2-}$ , and  $\text{C}_6\text{S}_6^{2-}$  as well as neutral species such as  $\text{C}_3\text{S}_3$ ,  $\text{C}_3\text{S}_8$ , and  $\text{C}_6\text{S}_{12}$ .

Two other compounds containing carbon and sulfur should be mentioned. The first of these is carbon monosulfide,  $\text{CS}$ . This compound has been reported to be obtained by the reaction of  $\text{CS}_2$  with ozone. The second compound is  $\text{COS}$  or, more correctly,  $\text{OCS}$ . It is prepared by the reaction



and it has melting and boiling points of  $-138.2$  and  $-50.2^\circ\text{C}$ , respectively. Unlike  $\text{CS}_2$ , neither  $\text{CS}$  nor  $\text{OCS}$  have significant industrial uses.

When  $\text{CS}_2$  is subjected to an electric discharge, one of the products is  $\text{C}_3\text{S}_2$ , a linear molecule having a structure similar to that of  $\text{C}_3\text{O}_2$  except for having sulfur atoms in the terminal positions.

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## PROBLEMS

1. Explain why living organisms maintain a relatively constant level of  $^{14}\text{C}$ .
2. The major industrial use of carbon is as a reducing agent. Write equations for three important processes that employ carbon as a reducing agent.
3. Write balanced equations to show the difference between methanides and acetylides in their reaction with water.
4. Show clearly why it is the carbon end of the  $\text{CO}$  molecule that binds to metals when carbon monoxide complexes are formed.
5. Consider a  $-1$  ion that contains one atom each of C, S, and P. Draw the correct structure and explain why any other arrangements of atoms are unlikely.
6. Draw the molecular orbital energy level diagram for the  $\text{C}_2^{2-}$  ion. Identify three species that are isoelectronic with  $\text{C}_2^{2-}$ .
7. Describe the process of casehardening of steel.
8. Using equations for the reactions, show how lime functions in mortar.
9. Predict some of the chemical characteristics of  $\text{CSe}_2$ .

10. Write complete, balanced equations to show the following processes.
  - (a) The reaction of  $\text{Ca}(\text{HCO}_3)_2$  with aqueous  $\text{NaOH}$ .
  - (b) The preparation of calcium cyanamide.
  - (c) The oxidation of potassium cyanide with  $\text{H}_2\text{O}_2$ .
  - (d) The reaction of methane with sulfur vapor.
  - (e) The reaction of acetylene with sodium hydroxide.
11. Draw the structure for the  $\text{C}_3\text{O}_2$  molecule. Identify all the symmetry elements it possesses and assign the point group (symmetry type).
12. Explain why  $\text{Co}(\text{CO})_4$  is not a stable molecule but solids containing the  $\text{Co}(\text{CO})_4^-$  ion are known.
13. Explain why the stable carbonyl of vanadium is not  $\text{V}(\text{CO})_6$  but is rather  $[\text{V}(\text{CO})_6]_2$ .
14. Would a solution of  $\text{NH}_4\text{CN}$  be acidic, basic, or neutral? For  $\text{NH}_3$ ,  $K_b = 1.8 \times 10^{-5}$  and for  $\text{HCN}$ ,  $K_a = 7.2 \times 10^{-10}$ . Write equations to help you arrive at your answer.
15. What would be the products of heating solid  $\text{BaCS}_3$ ?
16. Draw resonance structures for the molecule containing one atom each of C, S, and Se. Estimate the contribution of each structure. What experimental techniques could you use to determine if the contributions you assign are reasonable?
17. The formation of  $\beta$ -graphite from  $\alpha$ -graphite is accompanied by an enthalpy change of about  $0.60 \text{ kJ mol}^{-1}$  whereas the conversion of  $\alpha$ -graphite to diamond has an enthalpy change of about  $1.90 \text{ kJ mol}^{-1}$ . Although neither enthalpy change is very large, it is much easier to transform  $\alpha$ -graphite to the  $\beta$  form than it is to convert it to diamond. Explain this difference.

# Silicon, Germanium, Tin, and Lead

This chapter is devoted to the chemistry of two of the oldest known elements, tin and lead, and two that are of much more recent discovery, silicon (1824) and germanium (1886). The elements in Group IVA show the trend toward more metallic character in progressing down in the group. Silicon is a nonmetal, tin and lead are metals, and germanium has some of the characteristics of both metals and nonmetals. These elements are all of considerable economic importance, but in vastly different ways.

## 12.1 THE ELEMENTS

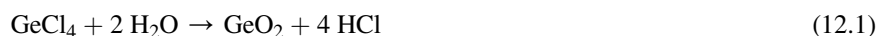
Tin and lead have been known since ancient times. *Cassiterite*,  $\text{SnO}_2$ , was mined in Britain and transported by sea to the Mediterranean area where copper was available. After reducing the  $\text{SnO}_2$  with charcoal to produce tin, the tin was alloyed with copper to make bronze as early as about 2500 BC. Consequently, tools and weapons made of bronze figured prominently in the period known as the Bronze Age (about 2500 to 1500 BC). At an early time, lead was found as native lead or as *galena*,  $\text{PbS}$ , that could be converted to the oxide by roasting the sulfide in air followed by reduction with carbon. As a result, tin and lead are among the elements known for many centuries. Of course, the reason that the metals Sn, Cu, Au, Ag, and Pb were available to the ancients is that either they were found uncombined (native) or they were easily reduced with charcoal (carbon). Today, the major sources of tin are Britain, Malaysia, Indonesia, and China.

Silicon was discovered by Berzelius in 1824. Although knowledge of the element itself is fairly recent, compounds of silicon have been known for thousands of years. For example, pottery, brick, ceramics, and glass are made of silicates that are naturally occurring materials. In addition to these uses of silicates, highly purified elemental silicon is used in the manufacture of integrated circuits or “chips” that are used in the electronics industry. It is also alloyed with iron to make *Duriron* (84.3% Fe, Si, 14.5%; Mn, 0.35%; and C, 0.85%), an alloy that resists attack by acids. This alloy is used in the manufacture of drain pipes and the cores of electric motors.

Silicon is an abundant element that makes up 23% of the earth’s crust, primarily in the form of silicate minerals and  $\text{SiO}_2$  (sand, quartz, etc.). The name silicon is derived from the Latin names *silex* and *silicis*, which refer to flint. The element is a brittle solid that has the diamond structure (density  $2.3 \text{ g cm}^{-3}$ ) and a gray luster. There is also an amorphous form of silicon that has a brown color.

In 2010, Geim and Novoselov received the Nobel Prize in Physics for experiments dealing with graphene, a compound composed of individual layers of carbon atoms removed from graphite. In 2012, it was reported that silicene, the silicon analog of graphene, had been synthesized by condensing a vapor of silicon atoms on a crystal of silver. The structure of silicene is similar to that of graphene except in silicene the layers are slightly buckled in a chair-type arrangement and are not planar. Because of this, silicene possesses a small band gap that, unlike graphene, allows silicene to function as a semiconductor. Unfortunately, silicene is unstable in air and degrades very rapidly.

Until 1871, the element having atomic number 32 was unknown. Based on atomic properties of the known elements, Mendeleev predicted that the element should resemble silicon, and he gave it the name *ekasilicon*, Es. Predicted properties were an atomic weight of 72 (the actual value is 72.59) and a density of  $5.5 \text{ g cm}^{-3}$  (the actual value is  $5.32 \text{ g cm}^{-3}$ ). In 1886, Winkler analyzed an ore, *argyrodite*, and found that it contained about 7% of an unidentified element. That element was germanium and its name comes from the Latin *Germania* for Germany. Currently, germanium is obtained primarily as a by-product from the separation of zinc from its ores. Treating the residue with concentrated HCl converts the germanium to  $\text{GeCl}_4$  that undergoes the hydrolysis reaction



Germanium dioxide is then reduced with hydrogen to obtain the element.





Elemental germanium is used primarily in the preparation of semiconductors in which it is combined with phosphorus, arsenic, or antimony to make n-type semiconductors or with gallium to make p-type semiconductors.

Tin is a soft, silvery metal with a very slight bluish color. Metallic tin has three forms that exist at different temperatures. At temperatures above 13.2 °C, the stable form has metallic properties, and it is known as *white* tin. When tin is heated above 161 °C, the white form is converted to a form known as *brittle* tin. This form does not behave as a metal, and it fractures when struck with a hammer. The form of tin that is stable below 13.2 °C is known as gray tin because it is easily crumbled to yield a gray powder. Even at temperatures below 13.2 °C, this transformation from the white form to the gray form is very slow unless the temperature is approximately –50 °C. The crumbling of gray tin (the low temperature form) is known historically as “tin disease” or “tin pest.”

The transformation of white tin to the gray form has some historical interest. In 1910–1912, Captain Robert Scott led a disastrous expedition to the South Pole. Scott’s party reached the South Pole only to find that an expedition led by Roald Amundsen had reached the pole a month earlier. Scott’s party had established a series of camps along their route where they had deposited stores of food and fuel. During their return trip, they arrived at one of their camps along their route only to discover that the metal containers holding fuel and supplies had failed. The metal containers had been sealed with tin and at the very low temperature in Antarctica the tin had changed into the gray form and crumbled. Leaving that camp, the party proceeded to the next camp on their route and found no fuel there for exactly the same reason. Scott’s party was too exhausted to continue the return to their base camp, and the entire party perished.

Tin and lead are often used together in a variety of useful alloys. Part of the reason for this is that the presence of a few percent of tin mixed with lead causes the alloy to be considerably harder than lead is alone. For example, common solder consists of about 50% Sn and 50% Pb. Type metal consists of about 82% Pb, 15% Sb, and 3% Sn. Pewter, an alloy used to make ornamental objects and vessels for food and beverages, consists of about 90% Sn with the remainder consisting of copper and antimony. Large quantities of tin are used as a coating for other metals, particularly steel, to retard corrosion. Alloys used as bearings often contain tin, antimony, and copper or aluminum. One such alloy is *babbit* that consists of 90% Sn, 7% Sb, and 3% Cu. Several other commonly encountered alloys have various specialty uses.

Lead has been used since perhaps 3000–4000 BC. Lead was mined by the Egyptians in the Sinai, and it was used for making glazes for pottery and vessels to hold liquids. The name *plumbum nigrum* was used by the Romans for lead, and the symbol Pb comes from the Latin name, *plumbum*. Many words such as plumbing, plumber, and plumb bob come from the name for the element.

The use of lead for dishes, roofs, and apparatus for collecting and holding water was widespread in Rome. Highly colored lead compounds were commonly used as cosmetics. At that time, lead (as well as tin) was obtained from Spain and England. Lead compounds are often found in minerals that also contain zinc compounds, as was the case with the mines of Laurium near Athens that were worked as early as 1200 BC.

Lead is almost always found combined, usually as the sulfide. However, lumps of lead have been found in areas where forest fires have occurred because some lead compound has been reduced under the ashes on the surface. Under these conditions, the sulfide is converted to the oxide,



Carbon (or carbon monoxide) resulting from incomplete combustion during the fire serves as the reducing agent.



These processes are essentially the same as those by which lead is obtained from its ore, and they could easily occur during forest fires. A number of other reactions are believed to occur during the reduction of PbS to produce lead. For example, the reaction of the oxide and the sulfide can be shown as follows:



In ancient times, the reduction process was carried out by heating PbO in a charcoal fire. Lead melts at 328 °C so molten lead can easily be separated. Missouri, Idaho, Utah, and Colorado are the largest producers of lead in the United States. Wisconsin and Illinois were formerly important producers as well. In addition to the United States, the largest producers of lead are Australia, Canada, and Russia.

In the United States, about 40% of the lead used is recovered from scrap. There are two types of scrap lead. The term “old scrap” applies to metal that has already been used in some manufactured object and the used article has been recycled

for resmelting. Lead in the form of pieces cut off, filings, turnings, etc., discarded during manufacturing processes is called “new scrap.” Automobile batteries constitute the largest source of old scrap lead.

Lead poisoning is still a problem today in some countries because lead and all of its compounds are toxic to humans. In certain parts of the world, lead compounds are still used as glazes and paints. It is believed that lead poisoning must have been a major problem for the Romans as the result of the widespread use of lead and lead-containing materials. Analysis of human remains from archaeological excavations in that region shows a high concentration of lead.

The Roman aristocracy had greater access to lead vessels and cosmetics containing lead compounds. It is believed that their life expectancy may have been as low as 25 years due to lead poisoning. In the body, lead accumulates in bones and the central nervous system. The production of hemoglobin is inhibited by lead because it binds to the enzymes that catalyze the reaction. High levels of lead cause anemia, kidney disfunction, and brain damage to occur, and the accumulation of lead interferes with proper development of the brain in children. Because of their toxicity, lead and its compounds are used much less today as paints or glazes than they were in earlier times.

Lead is a dense ( $11.4 \text{ g cm}^{-3}$ ) silvery metal having a face-centered cubic closest packed structure (see Chapter 4). It is widely used in the lead storage batteries in automobiles where the plates are made of an alloy containing about 88–93% Pb and 7–12% Sb. Many useful alloys of lead contain antimony because it produces an alloy that is stronger and harder than pure lead. We have already mentioned the uses of lead in solder, pewter, and type metal. It is also used in small arms ammunition. However, its use in shot shells for hunting waterfowl has been prohibited owing to its toxicity when ingested by these birds because they are largely bottom feeders. Steel shot is now required for this type of hunting and more recently shells containing bismuth shot have been approved.

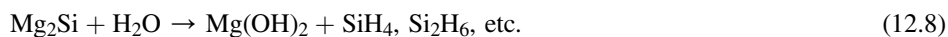
Lead compounds have long been used in pigments in paints, but this use has largely been discontinued. In white paint,  $\text{TiO}_2$  is now used as the pigment. About 75% of the lead used in the US is used in production of other chemicals. The use of tetraethyllead,  $\text{Pb}(\text{C}_2\text{H}_5)_4$ , as an antiknock agent in motor fuels formerly was responsible for the consumption of a large amount of lead.

## 12.2 HYDRIDES OF THE GROUP IVA ELEMENTS

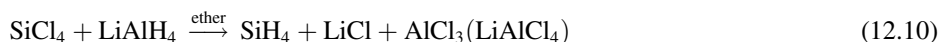
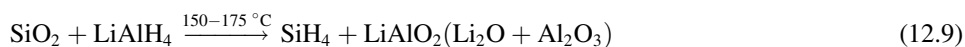
The elements Si, Ge, Sn, and Pb all exhibit the oxidation states of +2 and +4. However, the +2 state for Si is rare. One reason for this is that  $\text{SiO}$  is not stable and the halides  $\text{SiF}_2$  and  $\text{SiCl}_2$  are polymeric solids. A few Ge(II) compounds are known, e.g.,  $\text{GeO}$ ,  $\text{GeS}$ , and  $\text{GeI}_2$ . The +2 and +4 oxidation states are about equally common for Sn and Pb. For example,  $\text{SnO}_2$  is the most common ore of Sn, and there are numerous compounds containing Sn(II) (stannous compounds). As we will see later, there are also numerous common compounds of both Pb(II) and Pb(IV).

The most common hydrides of the Group IVA elements (represented as a group by E) are those having the formulas  $\text{EH}_4$ . They are covalent or volatile hydrides, and general characteristics of these compounds were discussed in Chapter 7. The hydrides  $\text{SiH}_4$ ,  $\text{GeH}_4$ ,  $\text{SnH}_4$ , and  $\text{PbH}_4$  decrease in stability in that order ( $\text{PbH}_4$  is very unstable). Such a trend is expected on the basis of the less effective overlap of the hydrogen 1s orbital with larger orbitals used by the atoms in Group IVA (see Chapter 3). In analogy to the hydrocarbons, the Group IVA hydrides,  $\text{EH}_4$ , are named as silane, germane, stannane, and plumbane when E represents Si, Ge, Sn, and Pb, respectively.

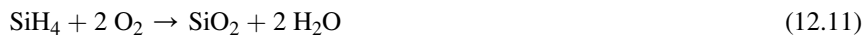
Alfred Stock prepared silicon hydrides by the means of the following reactions:



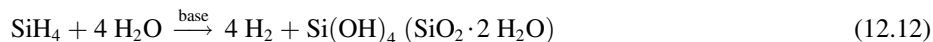
Silane can also be prepared by the reaction of  $\text{SiO}_2$  or  $\text{SiCl}_4$  with  $\text{LiAlH}_4$  as represented by the equations



Silane and disilane ( $\text{Si}_2\text{H}_6$ ) are the only stable hydrides of silicon because the higher members of the series decompose to produce  $\text{SiH}_4$ ,  $\text{Si}_2\text{H}_6$ , and  $\text{H}_2$ . Silicon hydrides are spontaneously flammable in air (see Chapter 7).



Undoubtedly, the driving force for this reaction is the extreme stability of  $\text{SiO}_2$  for which the heat of formation is  $-828 \text{ kJ mol}^{-1}$ . Although  $\text{SiH}_4$  and  $\text{Si}_2\text{H}_6$  are stable in water and dilute acids, they react with water in basic solutions as shown by the equation



Germane is prepared by the reaction of  $\text{GeO}_2$  with  $\text{LiAlH}_4$ .



The higher germanes are obtained by electric discharge through  $\text{GeH}_4$ . The flammability of the germanes is less than that of the silanes, and they do not hydrolyze. Stannane can be prepared by the reaction

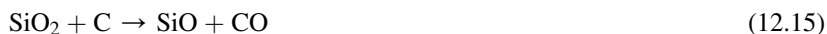


## 12.3 OXIDES OF THE GROUP IVA ELEMENTS

Silicon forms both  $\text{SiO}$  and  $\text{SiO}_2$ , the latter existing with Si atoms surrounded tetrahedrally by four oxygen atoms. Germanium, tin, and lead form both monoxides and dioxides, but the dioxides are structurally much simpler than those of  $\text{SiO}_2$ . From the outset, it should be recognized that it is difficult to distinguish between the oxide, the hydrous oxide (or hydrated oxide), and the hydroxide. For example,  $\text{GeO}$ ,  $\text{GeO} \cdot x\text{H}_2\text{O}$ , and  $\text{Ge}(\text{OH})_2$  (the formula for the hydroxide is also equivalent to  $\text{GeO} \cdot \text{H}_2\text{O}$ ) all exist, sometimes in equilibria or in mixtures. A similar situation is found for  $\text{SnO}_2$ ,  $\text{SnO}_2 \cdot x\text{H}_2\text{O}$ , and  $\text{Sn}(\text{OH})_4$  (that can also be written as  $\text{H}_4\text{SnO}_4$ ). The oxides of the Group IVA elements will be considered according to the usual oxidation states of the elements.

### 12.3.1 The +2 Oxides

The unstable oxide  $\text{SiO}$  can be regarded as the silicon analog of carbon monoxide, the product obtained when carbon reacts with a deficiency of oxygen. However, it is also possible to obtain  $\text{SiO}$  by the reaction of  $\text{SiO}_2$  with carbon.



The monoxide is unstable because  $\text{SiO}_2$  is very stable. As a result, the disproportionation reaction



is energetically favorable.

Germanium(II) oxide is a black powder that can be prepared by the reaction



When heated to approximately  $500^\circ\text{C}$ ,  $\text{GeO}$  disproportionates to  $\text{Ge}$  and  $\text{GeO}_2$ .



The oxide of  $\text{Sn}(\text{II})$  can be prepared by dehydrating  $\text{Sn}(\text{OH})_2$ ,



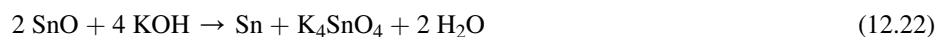
The hydroxide is produced when  $\text{SnCl}_2$  undergoes the hydrolysis reaction



When  $\text{SnO}$  is heated with a strong base, disproportionation occurs to produce an oxyanion of tin.

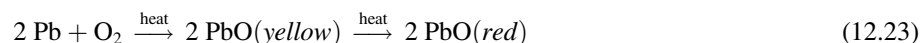


This reaction can also take place with a different stoichiometry as is illustrated by the equation



However, an oxyanion containing Sn(IV) is produced in either case, and both  $\text{SnO}_3^{2-}$  and  $\text{SnO}_4^{4-}$  are referred to as *stannates*.

Two forms of PbO known as *litharge* (red) and *massicot* (yellow) are obtained by reacting lead with oxygen.



These forms of PbO have been used as pigments for many years. As was mentioned earlier, there is also difficulty in distinguishing between the oxide (PbO), the hydroxide ( $\text{Pb}(\text{OH})_2$ ), and the hydrous oxide ( $\text{PbO} \cdot x\text{H}_2\text{O}$ ).

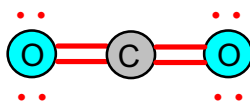
### 12.3.2 The +4 Oxides

There is a much more extensive chemistry of the +4 oxides of the Group IVA elements than there is for the +2 oxides. In general, the  $\text{EO}_2$  compounds are acidic or amphoteric oxides, and they show this characteristic by forming oxyanions. This type of behavior has also been illustrated for  $\text{CO}_2$  by the reaction



Similar behavior for the remaining elements in Group IVA leads to the formation of silicates, stannates, etc.

Although the oxides of all the Group IVA elements are of some interest, the chemistry of  $\text{SiO}_2$  is by far the most complex. Unlike  $\text{CO}_2$  where double bonding results in the molecule having the structure



the Si—O bond is quite strong, and there is very little tendency for  $\text{Si}=\text{O}$  double bonds to form. Therefore, each Si must form four bonds to oxygen atoms and each O must form two bonds to silicon atoms to satisfy the valences of +4 and  $-2$ , respectively (see Chapter 4). As a result, Si atoms occupy sites in which they are surrounded tetrahedrally by four oxygen atoms and the oxygen atoms form bridges between silicon atoms. Such a situation can result in a great deal of polymorphism, and over 20 forms of  $\text{SiO}_2$  exist. *Quartz*, *tridymite*, and *crystalobolite* are three forms of  $\text{SiO}_2$  and each of these exists in  $\alpha$  and  $\beta$  forms. There is interconversion between some of the forms upon heating.

Silicon dioxide is found as sand, flint, quartz, agate, and other common materials. The compound melts at  $1710^\circ\text{C}$ , but it softens at temperatures below that when some of the Si—O—Si bridges are broken. Molten  $\text{SiO}_2$  is a thick liquid that cools to give a glass in which only part of the Si—O bonds have been disrupted.

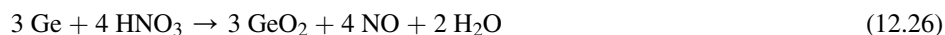
When it is struck or pressed, quartz generates an electric current. Materials having this property are known as *piezoelectric* materials. If an external voltage is applied across the crystal, the crystal undergoes vibrations that are in resonance with the alternating current frequency. This type of behavior is the basis for quartz being used as a timing device in watches or in crystals used to establish radio frequencies.

Melting quartz and allowing it to cool results in a type of “quartz glass” that is considerably different from most types of glass. It is transparent to electromagnetic radiation over a wide range of wave lengths, and, as a result, it is used in making optical devices. Because of the importance of glass in today’s economy, it will be discussed later in a separate section of this chapter.

Germanium dioxide can be prepared by the direct reaction of the elements,



or by the oxidation of Ge with nitric acid as represented in the equation



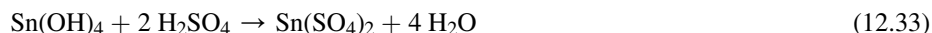
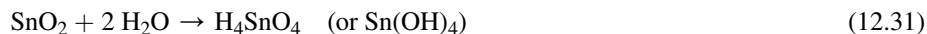
It is also produced by the reaction of  $\text{GeCl}_4$  with hydroxides,



When dissolved in water,  $\text{GeO}_2$  produces a solution that is a weak acid because of the following reactions and the slight dissociation of the acids produced.



*Cassiterite* has the rutile structure (see Chapter 4), and it is the naturally occurring mineral form of  $\text{SnO}_2$ . As shown by the following reactions, this oxide is amphoteric:

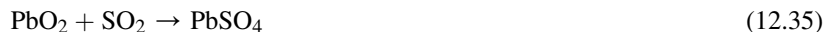


Both  $\text{SnO}_4^{4-}$  and  $\text{SnO}_3^{2-}$  are *stannates* because they contain Sn(IV). The former is sometimes known as *orthostannate* and the latter as *metastannate* in the same way that  $\text{PO}_4^{3-}$  and  $\text{PO}_3^-$  are known as the orthophosphate and metaphosphate ions, respectively (see Chapter 14).

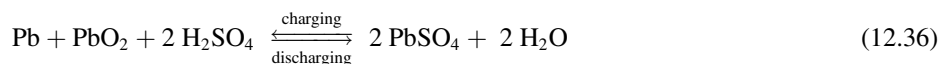
Lead dioxide also has the rutile structure (see Chapter 4), and it can be prepared by the oxidation of  $\text{PbO}$  with  $\text{NaOCl}$  as shown in the following equation:



Lead dioxide is a strong oxidizing agent, and it will oxidize  $\text{SO}_2$  to sulfate.



$\text{PbO}_2$  is the oxidizing agent in the lead storage battery where the reducing agent is metallic lead. The overall reaction for the battery is



The battery contains cells in which the electrodes are lead and spongy lead impregnated with  $\text{PbO}_2$ . In these cells,  $\text{H}_2\text{SO}_4$  is the electrolyte, and each cell produces approximately 2.0 V so that six cells linked in series result in a 12-volt battery.

In addition to  $\text{PbO}$  and  $\text{PbO}_2$ , lead forms two other oxides that are rather well known. The first is  $\text{Pb}_2\text{O}_3$ , sometimes written as  $\text{PbO} \cdot \text{PbO}_2$ , a compound that is more correctly considered to be  $\text{PbPbO}_3$  showing that it contains Pb(II) and Pb(IV). The second is known as red lead,  $\text{Pb}_3\text{O}_4$ , and it is used in making flint glass (lead crystal) and as a red pigment. It results from the heating of litharge ( $\text{PbO}$ ) in air at 400 °C.



The formula for red lead is sometimes written as  $2 \text{PbO} \cdot \text{PbO}_2$  or  $\text{Pb}_2\text{PbO}_4$ .

### 12.3.3 Glass

Sand (silica or silicon dioxide) melts at 1710 °C, and sodium carbonate melts at 851 °C. These ingredients along with lime,  $\text{CaO}$ , make up a common type of glass known as *soda-lime glass* because it contains soda, lime, and silicon dioxide ( $\text{Na}_2\text{CO}_3$ ,  $\text{CaO}$ , and  $\text{SiO}_2$ ) with the composition being about 25%, 12.5%, and 62.5%, respectively.

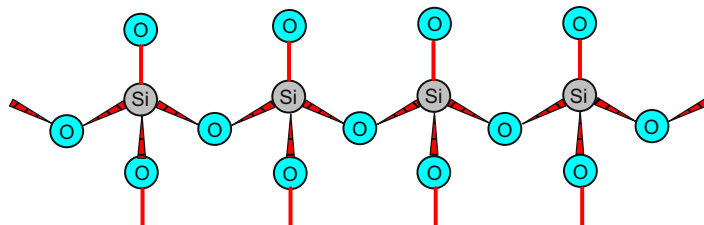
There is no way to tell exactly when and how glass making was discovered, but archaeological studies of tombs in Egypt show that glass objects were made as early as about 5000 years ago. It was made by heating a mixture of sand ( $\text{SiO}_2$ ), soda or sodium carbonate, ( $\text{Na}_2\text{CO}_3$ ), and limestone ( $\text{CaCO}_3$ ). When the mixture is heated, the limestone decomposes to produce lime and carbon dioxide,



When the molten mixture cools, it produces a transparent, rigid mass that softens when it is reheated. While it is hot and soft, the glass can be shaped by blowing, rolling, or molding it into the desired shape. The glass used most often in windows, sometimes called flat glass, is of the soda-lime type described above.

The shaping of blown glass is believed to date from the first century BC. Molten glass can be collected on the end of a blow pipe and blown into shape while rotating the pipe. The molten mass can be encased in a mold so that the blown glass is given the shape of the mold. The shaped glass object should be placed in a furnace and allowed to cool slowly in order to anneal it.

In glass, the silicon and oxygen atoms are arranged in tetrahedral  $\text{SiO}_4$  units. Oxygen atoms bridging between two silicon atoms give chains and network structures in three dimensions. This arrangement can be represented as



Heating glass causes the disruption of some of the bonds resulting in a less rigid mass and softening of the glass. Not all of the bonds are broken so the glass does not become a mobile liquid. Cooling the glass allows the formation of a network so that the glass becomes rigid, but not crystalline. However, if glass is cooled too slowly, a large fraction of the  $-\text{Si}-\text{O}-\text{Si}-\text{O}-$  linkages reform making the structure more closely resemble that of crystalline silicates, and the product is brittle.

Sheets of glass having very flat surfaces are obtained by placing molten glass on the surface of molten tin in a large, shallow container. The molten, dense metal forms a very flat surface and the molten glass floats on the surface of the metal. When it is allowed to cool, flat sheets are obtained.

Adding  $\text{B}_2\text{O}_3$  to sand and soda produces a type of glass known as *borosilicate glass*. An important characteristic of this type of glass is that it is heat resistant and does not shatter when heated or cooled quickly. Adding lead oxides ( $\text{PbO}$  and  $\text{Pb}_3\text{O}_4$ ) instead of lime results in glass known as *lead crystal* and *flint glass*. Because it is very dense and highly refractive to light, this type of glass is frequently used in making ornamental glass objects.

The addition of many other substances to sand, soda, and lime produces specialty glasses having different colors. Some of the materials that can be added and the colors of glasses that result are shown below.

Additive	Color of glass produced
$\text{CoO}$	Blue
$\text{Cu}_2\text{O}$	Red, green, or blue
$\text{SnO}_2$	Opaque
$\text{Fe}_2\text{O}_3$	Yellow
$\text{FeO}$	Green
$\text{Au}$	Red, blue, or purple
$\text{CaF}_2$	Milky white

Iron compounds (particularly oxides) are very widespread in nature, and they are often present in small amounts in sand and limestone. Inexpensive glass sometimes has a slightly greenish color that can be avoided if raw materials of high purity are used. Very old glass bottles frequently have a greenish tint because of this condition.

Glass is a versatile material that is made from inexpensive materials that are widely distributed in nature, and it is recyclable. It can be shaped by a variety of techniques to produce articles of almost any configuration. Glass objects are durable, transparent, and easily cleaned, and they can be given almost any color by using the appropriate additives. Glass is a good electrical insulator and unreactive toward most chemicals. Specialty glasses are of enormous importance in the construction of optical devices and scientific equipment. It is no wonder that glass is used in enormous quantities with the annual US production of flat and container glass alone amounting to million of tons.

In recent years, a large and increasing amount of glass has been used in the construction industry. Hardened or tempered glass has many desirable qualities as a construction material. It is structurally strong, it lasts indefinitely with little

maintenance, and it is made from inexpensive and widely available materials. There is little doubt that this trend will continue and there will be increased use of glass in this way in the future especially with metals becoming more scarce and expensive. Glass has been in use for perhaps 5000 years, and its importance has grown throughout that time.

## 12.4 SILICATES

As was mentioned in the section dealing with oxides,  $\text{SiO}_2$  exists in a number of forms. When molten magma hardens in the earth's crust, a wide variety of silicate minerals are formed. Many of these can formally be considered as being derived from the interaction of  $\text{SiO}_2$  with metal oxides where the acidic nature of  $\text{SiO}_2$  results in the reaction



Some of the common combinations of oxides and the resulting minerals are illustrated in the following table:

Oxides combined	Mineral equivalent
$\text{CaO} + \text{TiO}_2 + \text{SiO}_2$	$\text{CaTiSiO}_5$ , <i>titanite</i>
$\frac{1}{2} \text{K}_2\text{O} + \frac{1}{2} \text{Al}_2\text{O}_3 + 3 \text{SiO}_2$	$\text{KAlSi}_3\text{O}_8$ , <i>orthoclase</i>
$2 \text{MgO} + \text{SiO}_2$	$\text{Mg}_2\text{SiO}_4$ , <i>forsterite</i>
$\text{BaO} + \text{TiO}_2 + 3 \text{SiO}_2$	$\text{BaTiSi}_3\text{O}_9$ , <i>benitoite</i>

Because of the way in which some minerals are regarded as combinations of oxides, it is not uncommon to find the composition of a silicate mineral given in terms of the oxides that formally constitute it. For example, *benitoite* is described as 36.3% BaO, 20.2%  $\text{TiO}_2$ , and 43.5%  $\text{SiO}_2$  in some books dealing with mineralogy.

When it is remembered that an oxide is easily converted to a hydroxide by the reaction with water,

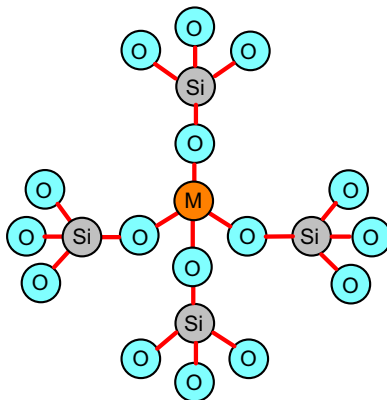


and that an oxide is easily converted to a carbonate by the reaction with carbon dioxide,



it is easy to understand how many minerals are interconverted under the action of water and carbon dioxide. Of course,  $\text{F}^-$  easily replaces  $\text{OH}^-$  in some structures so even more possibilities exist. Moreover,  $\text{H}_2\text{O}$  may be incorporated in the structures. The subject of silicate minerals is so vast that an entire book could easily be devoted to it. Silicate chemistry will be described only briefly here, and the references at the end of the chapter should be consulted for more detailed coverage.

Just as  $\text{PO}_4^{3-}$  is known as the orthophosphate ion, the tetrahedral  $\text{SiO}_4^{4-}$  ion is called *orthosilicate*. This tetrahedral monomeric unit is found in several minerals, e.g., *phenacite*,  $\text{Be}_2\text{SiO}_4$ , and *willemite*,  $\text{Zn}_2\text{SiO}_4$ . In both of these minerals, the metal is surrounded by four oxygen atoms that are bound in  $\text{SiO}_4$  units arranged tetrahedrally around the metal.



However, the number of  $\text{SiO}_4^{4-}$  units that function as electron pair donors through the oxygen atoms can vary. For example, in  $\text{Mg}_2\text{SiO}_4$  and  $\text{Fe}_2\text{SiO}_4$  (forms of *olivine*), the coordination number of the metals is six but *zircon*,  $\text{ZrSiO}_4$ , has a structure in which the coordination number of Zr is eight.

The complex structures of silicates are the result of the numerous ways in which the  $\text{SiO}_4$  tetrahedra can be arranged. An interesting shorthand that has developed to show the structures of the silicate ions is illustrated in [Figure 12.1](#)

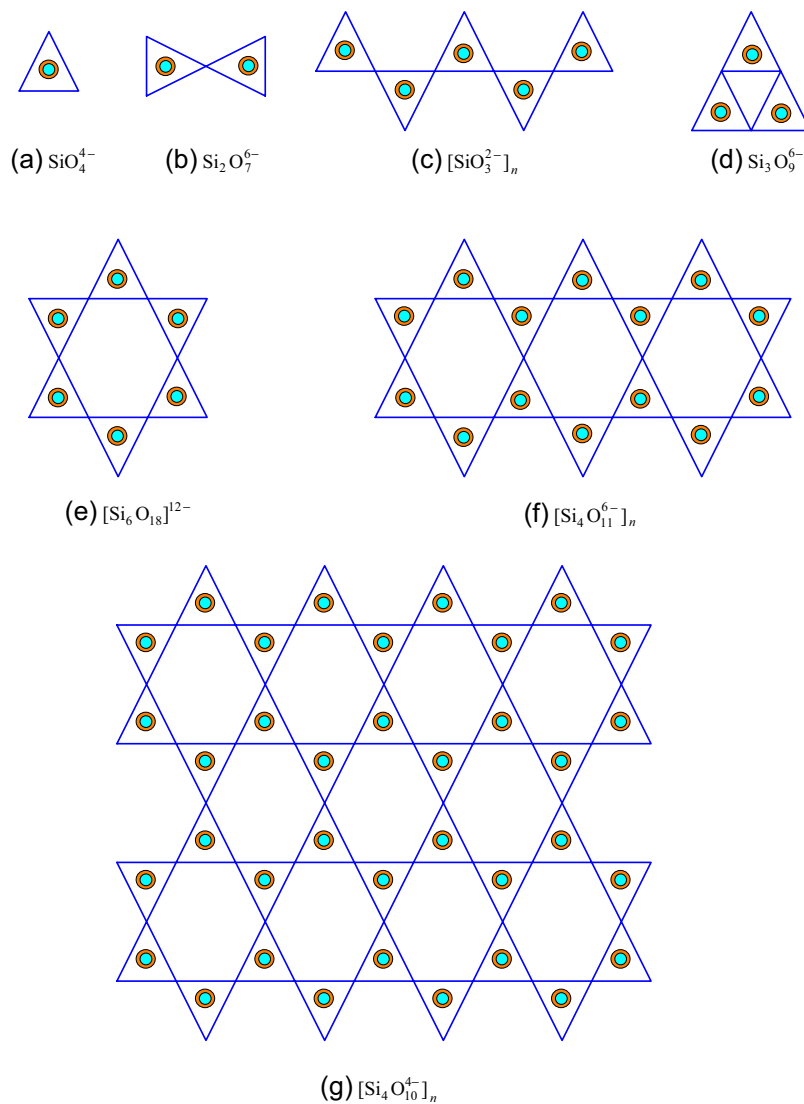
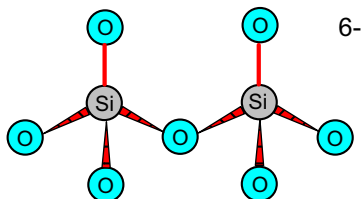


FIGURE 12.1 The structures of the silicates.

(the tetrahedral  $\text{SiO}_4^{4-}$  ion is shown as Figure 12.1(a)). In these structures, it is presumed that a “top view” of a tetrahedron is shown so that the blue circle represents the “top” oxygen atom, the orange ring represents the Si atom directly below the oxygen atom, and the triangle represents the triangular base of the tetrahedron formed by the other three oxygen atoms. The structures of many silicates are then made up by sharing one or more corners of the tetrahedra where oxygen atoms form bridges as shown in Figure 12.1(b)–(g).

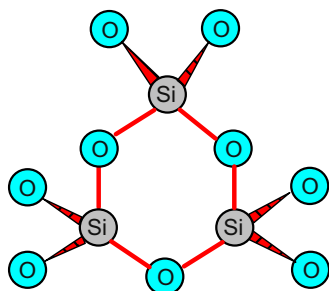
Two tetrahedral units sharing one corner result in the structure known as *pyrosilicate* (see Figure 12.1(b)).



This structure is similar to those of the isoelectronic species  $\text{P}_2\text{O}_7^{4-}$ ,  $\text{S}_2\text{O}_7^{2-}$ , and  $\text{Cl}_2\text{O}_7$  (the pyrophosphate and pyrosulfate ions, and dichlorine heptoxide). The minerals *thortveitite*,  $\text{Sc}_2\text{Si}_2\text{O}_7$ , and *hemimorphite*,  $\text{Zn}_4(\text{OH})_2\text{Si}_2\text{O}_7$ , contain the  $\text{Si}_2\text{O}_7^{6-}$  unit which has the structure shown above.



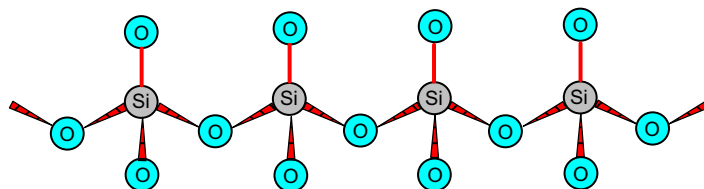
Three  $\text{SiO}_3^{2-}$  (metasilicate) ions can be linked by sharing two oxygen atoms to give a cyclic structure,  $\text{Si}_3\text{O}_9^{6-}$ , that consists of three  $\text{SiO}_4$  tetrahedra each sharing two corners. Minerals containing this structure belong to the class known as the metasilicates. The ring structure of the  $\text{Si}_3\text{O}_9^{6-}$  can be illustrated as shown below (see also Figure 12.1(d)).



This structure is analogous to  $\text{P}_3\text{O}_9^{3-}$  and  $\text{S}_3\text{O}_9$  (a form of solid  $\text{SO}_3$  having the structure  $(\text{SO}_3)_3$ ). A substance that contains the  $\text{Si}_3\text{O}_9^{6-}$  ion is the mineral *benitoite*,  $\text{BaTiSi}_3\text{O}_9$ .

Linking six  $\text{SiO}_3^{2-}$  ions in a ring gives the  $\text{Si}_6\text{O}_{18}^{12-}$  ion that is present in *beryl*,  $\text{Be}_3\text{Al}_2\text{Si}_6\text{O}_{18}$ . In this case, the  $\text{Si}_6\text{O}_{18}^{12-}$  ion contains a 12-membered ring having alternating Si and O atoms. The structure is represented in Figure 12.1(e).

Silicate chains are of two general types. Long chains of  $\text{SiO}_4$  tetrahedra are present in minerals known as *pyroxenes* that can be considered as a repeating pattern of  $\text{SiO}_3$  units giving rise to a structure such as that shown below having the formula  $(\text{SiO}_3^{2-})_n$  (see also Figure 12.1(c)).



The individual chains are held together by sharing metal ions located between them. Examples of this type include *diopside* ( $\text{CaMgSi}_2\text{O}_6$ ), *hedenbergite* ( $\text{Ca}(\text{Fe},\text{Mg})\text{Si}_2\text{O}_6$ ), and *spodumene* ( $\text{LiAlSi}_2\text{O}_6$ ).

The *amphiboles* (shown in Figure 12.1(f)) contain double chains consisting of  $\text{SiO}_4$  tetrahedra that are represented as  $(\text{Si}_4\text{O}_{11}^{6-})_n$  units. These chains are also held together by being bound to the metal ions. However, in the case of the amphiboles, half of the silicon atoms share two oxygen atoms between them whereas three oxygen atoms are shared by the other half of the silicon atoms (see Figure 12.1(f)). Minerals of this type are usually considered to be formed from silicates when water is present under high temperature and pressure. Examples of amphiboles include *tremolite*,  $\text{Ca}_2\text{Mg}_5\text{Si}_8\text{O}_{22}(\text{OH})_2$ , and *hornblende*,  $\text{CaNa}(\text{Mg},\text{Fe})_4(\text{Al},\text{Fe},\text{Ti})_3\text{Si}_6\text{O}_{22}(\text{O},\text{OH})_2$ , among others. Additional chains are bound together when three oxygen atoms are shared by all of the silicon atoms so that a sheet structure characteristic of the *micas*. Bonding between the sheets is rather weak and the parallel sheets can easily be separated. Examples of micas are *muscovite*,  $\text{KAl}_3\text{Si}_3\text{O}_{10}(\text{OH})_2$ , *biotite*,  $\text{K}(\text{Mg},\text{Fe})_3\text{AlSi}_3\text{O}_{10}(\text{OH})_2$ , and *lepidolite*,  $\text{K}_2\text{Li}_3\text{Al}_4\text{Si}_7\text{O}_{21}(\text{OH},\text{F})_3$ . Being a rather common mineral, muscovite was formerly used in electrical insulators and, as a result of it forming transparent sheets, small windows in stoves. It has also been used as a dry lubricant.

The *feldspars* are derived from structures consisting of  $(\text{SiO}_2)_n$  in which each Si is surrounded by four O atoms and each O is surrounded by two Si atoms. To show the effect of substitution of metal ions for  $\text{Si}^{4+}$ , consider  $(\text{SiO}_2)_4$  or  $\text{Si}_4\text{O}_8$ . Substituting one  $\text{Al}^{3+}$  in place of an  $\text{Si}^{4+}$  produces  $\text{AlSi}_3\text{O}_8^-$ . Replacement of  $\text{Si}^{4+}$  in some  $\text{SiO}_4$  units by  $\text{Al}^{3+}$  thus opens the possibility for a +1 ion to be present so that a +4 ion is being replaced by the combination of a +3 and a +1 ion. In the example above, we have replaced one-fourth of the  $\text{Si}^{4+}$  by  $\text{Al}^{3+}$ . If we replace two  $\text{Si}^{4+}$  with  $\text{Al}^{3+}$ , we obtain  $\text{Al}_2\text{Si}_2\text{O}_8^{2-}$  and the other ion must be a positive ion having a +2 charge. That ion is frequently  $\text{Ca}^{2+}$ . The feldspars are aluminum silicates (or *aluminosilicates*) that contain  $\text{Na}^+$  or  $\text{K}^+$  as the +1 ion. For example, *orthoclase* and *microcline* are two forms of  $\text{KAlSi}_3\text{O}_8$  whereas *sanadine* (represented by the formula  $(\text{Na},\text{K})\text{AlSi}_3\text{O}_8$ ) has up to half of the potassium replaced by sodium. Showing the sodium and potassium as  $(\text{Na},\text{K})$  indicates that the mineral may contain either Na or K or a mixture of the two as long as the *total* number of ions provides for electrical neutrality. Partial substitution of metal ions in silicate minerals is quite common, and in some cases a +3 ion is replaced by the combination of +1 and +2 ions. This variability

in composition is illustrated by the *plagioclase* series shown in the following table that constitutes a group of closely related materials resulting from ionic substitution.

Name	Composition	Density, g cm <sup>-3</sup>
<i>Albite</i>	NaAlSi <sub>3</sub> O <sub>8</sub>	2.63
<i>Oligoclase</i>	(Na,Ca)AlSi <sub>2</sub> O <sub>8</sub>	2.65
<i>Andersine</i>	(Na,Ca)AlSi <sub>2</sub> O <sub>8</sub>	2.68
<i>Labradorite</i>	(Na,Ca)AlSi <sub>2</sub> O <sub>8</sub>	2.71
<i>Bytownite</i>	(Na,Ca)AlSi <sub>2</sub> O <sub>8</sub>	2.74
<i>Andorite</i>	CaAl <sub>2</sub> Si <sub>2</sub> O <sub>8</sub>	2.76

All of these minerals are used in ceramics and refractories, and *labradorite* is sometimes used as a building stone.

An important aluminosilicate formed from feldspar is *kaolin*, Al<sub>2</sub>Si<sub>2</sub>O<sub>5</sub>(OH)<sub>4</sub>, (sometimes described as having the composition 39.5% Al<sub>2</sub>O<sub>3</sub>, 46.5% SiO<sub>2</sub>, and 14.0% H<sub>2</sub>O). It is widely used in making ceramics, and in high purity it is used as a clay for making china. Closely related to kaolin is *orthoclase*, KAlSi<sub>3</sub>O<sub>8</sub>, a substance that is used in ceramics and certain types of glass. Another useful aluminosilicate is *leucite*, KAlSi<sub>2</sub>O<sub>6</sub>, which has been used as a source of potassium in fertilizer.

## 12.5 ZEOLITES

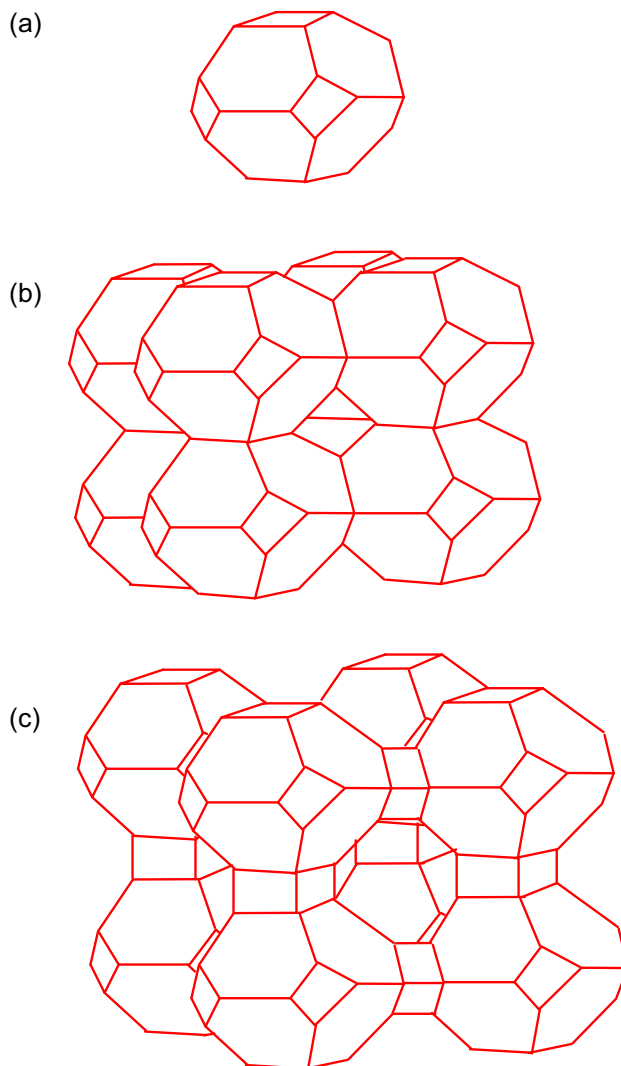
Minerals have been formed in a variety of environments where the temperature may have varied widely. Hydrous silicates known as zeolites are secondary minerals that form in igneous rocks. The name *zeolite* comes from the Greek *zeo*, meaning to boil, and *lithos*, meaning rock. The loss of water when zeolites are heated resulted in those terms being combined in the name that was given to them by Baron Axel Cronstedt in 1756. Zeolites are aluminosilicates that have large anions containing cavities and channels. More than 30 naturally occurring zeolites are known, but several times that number have been synthesized. Many of these have variable composition because sodium and calcium can be interchanged in the structures. Because of this, calcium ions in hard water are removed by the use of water softeners containing a zeolite that is high in sodium content. The zeolite saturated with Ca ions can be renewed by washing it with a concentrated salt solution whereby the calcium ions are replaced by sodium ions and the zeolite is again able to function effectively. In addition to their use in ion exchange processes (for these uses they have largely been replaced by ion exchange resins), zeolites are also used as molecular sieves and as catalysts for certain processes.

Zeolites have a composition that can be described by the general formula M<sub>a/z</sub>[(AlO<sub>2</sub>)<sub>a</sub>(SiO<sub>2</sub>)<sub>b</sub>]·xH<sub>2</sub>O, where M is a cation of charge z. Because the SiO<sub>2</sub> units and the H<sub>2</sub>O molecules are uncharged, the number of these constituents does not affect the stoichiometry required to balance the AlO<sub>2</sub><sup>-</sup> charges. If M is a +1 ion such as Na<sup>+</sup>, then each AlO<sub>2</sub><sup>-</sup> requires one cation. If M is a +2 ion such as Ba<sup>2+</sup>, one cation balances the charge on two AlO<sub>2</sub><sup>-</sup> ions. Thus, the number of M ions is related to the number of AlO<sub>2</sub><sup>-</sup> ions (a) by the ratio a/z:a. A few naturally occurring zeolites are listed below.

<i>Analcime</i>	NaAlSi <sub>2</sub> O <sub>6</sub> ·H <sub>2</sub> O
<i>Edingtonite</i>	BaAl <sub>2</sub> Si <sub>3</sub> O <sub>10</sub> ·4 H <sub>2</sub> O
<i>Cordierite</i>	(Mg,Fe) <sub>2</sub> Mg <sub>2</sub> Al <sub>4</sub> Si <sub>5</sub> O <sub>18</sub>
<i>Stilbite</i>	(Ca,Na) <sub>3</sub> Al <sub>5</sub> (Al,Si) <sub>14</sub> O <sub>40</sub> ·15 H <sub>2</sub> O
<i>Chabazite</i>	(Ca,Na,K) <sub>7</sub> Al <sub>12</sub> (Al,Si) <sub>2</sub> Si <sub>26</sub> O <sub>80</sub> ·40 H <sub>2</sub> O
<i>Sodalite</i>	Na <sub>2</sub> Al <sub>2</sub> Si <sub>3</sub> O <sub>10</sub> ·2 H <sub>2</sub> O

In the framework structures of zeolites, the basic unit is the tetrahedron with the composition SiO<sub>4</sub> or AlO<sub>4</sub>. In each case where Al<sup>3+</sup> replaces Si<sup>4+</sup>, a positive charge from another cation is required. Many of the structures contain the Si<sub>6</sub>O<sub>18</sub><sup>12-</sup> ion that has a ring structure with some of the Si atoms replaced by Al atoms (see Figure 12.1(e)). This ring is represented as a hexagonal face because it contains six silicon atoms with oxygen atoms surrounding each silicon atom, but they are not shown. These hexagons can be joined to give a structure that can be represented as shown in Figure 12.2(a).

This is the type of unit present in *sodalite* (also sometimes referred to as *natronite*), and it is known as a β-cage. The structure contains rings that are referred to as 6-rings and 4-rings because of the numbers of nonoxygen atoms in the rings. It should be noted that although the edges of the rings are drawn as straight lines, they represent Si—O—Si and Al—O—Si



**FIGURE 12.2** The structures of (a) the  $\beta$ -cage of a single sodalite unit, (b) the mineral sodalite showing eight  $\beta$ -cages, and (c) zeolite-A.

bridges that are not linear. Eight sodalite units can be joined in a cubic pattern as shown in [Figure 12.2\(b\)](#) with the units being joined by sharing faces composed of 4-membered rings. Such a structure has channels between the sodalite units on the corners of the cube. If the eight sodalite units are joined by bridging oxygen atoms at each corner of the 4-rings, the resulting structure is that of *zeolite-A* (shown in [Figure 12.2\(c\)](#)), a well-known synthetic zeolite having the formula  $\text{Na}_{12}(\text{AlO}_2)_{12}(\text{SiO}_2)_{12} \cdot 27 \text{H}_2\text{O}$  or  $\text{Na}_{12}\text{Al}_{12}\text{Si}_{12}\text{O}_{48} \cdot 27 \text{H}_2\text{O}$ . In this structure, half of the Si atoms have been replaced by Al atoms. Although the structures of many zeolites are known, they will not be discussed in more detail here.

Some zeolites are useful catalysts as a result of their having very large surface areas. Their channels and cavities can differentiate between molecules on the basis of their ability to migrate to and from active sites on the catalyst. In some cases, the  $\text{Na}^+$  ions are replaced by  $\text{H}^+$  ions that attach to oxygen atoms to give  $-\text{OH}$  groups. These sites are referred to as *Brønsted sites* because they can function as proton donors. *Lewis acid sites* occur when the protonated zeolite is heated so that  $\text{H}_2\text{O}$  is driven off, leaving Al that is coordinatively unsaturated. An Al atom having only three bonds to it is electron deficient so it can function as a site that capable of behaving as a Lewis acid.

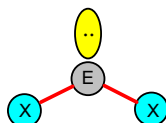
Zeolites containing multiply charged cations are effective catalysts for “cracking” hydrocarbons in petroleum. The substitution of  $\text{La}^{3+}$  or  $\text{Ce}^{3+}$  for  $\text{Na}^+$  produces zeolites in which there are regions of high charge density that can induce reactions of C–H bonds. The reactions produce a more highly branched structure, and such hydrocarbons give fuels with higher octane numbers. These catalytic processes represent the major ones used to convert a large fraction of petroleum into motor fuels. Processes have also been developed for production of motor fuels from methanol and for production of xylenes.

The chemistry of zeolites has considerable commercial application. However, the details of this important topic are beyond the scope of this book. As has been shown, the silicates constitute an enormous range of materials, and additional information on the various phases of silicate chemistry is available in the references listed at the end of this chapter.

## 12.6 HALIDES OF THE GROUP IV ELEMENTS

Two well-defined series of halogen compounds of the Group IVA elements (represented as E) having the formulas  $EX_2$  and  $EX_4$  are known. In addition, a few compounds having the formula  $E_2X_6$  are also known, especially for silicon ( $Si_2F_6$ ,  $Si_2Cl_6$ ,  $Si_2Br_6$ , and  $Si_2I_6$ ) in keeping with the greater tendency of Si atoms to bond with themselves. It is a general trend that when an element forms compounds in which it can have different oxidation states, the element behaves *less* like a metal the *higher* its oxidation state. The halogen compounds of the elements in the +2 oxidation state,  $EX_2$ , tend to be more like ionic salts whereas those containing the elements in the +4 oxidation state tend to be more covalent. Thus, the  $EX_2$  compounds tend to have higher melting and boiling points than do those of the  $EX_4$  series. Physical data reported in the literature for  $EX_2$  and  $EX_4$  compounds show considerable variation in some cases. Table 12.1 shows melting points and boiling points for the  $EX_2$  and  $EX_4$  compounds that have been well characterized.

With three pairs of electrons surrounding the central atom, the monomeric molecules are expected to have the nonlinear structure shown below (for example, the bond angle is  $101^\circ$  in  $SiF_2$ ), and, as a result, they are polar.



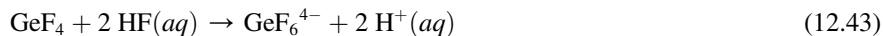
However, the  $EX_4$  molecules have tetrahedral structures and they are nonpolar. As a consequence, the compounds having the formula  $EX_4$  are more soluble in organic solvents than are those of the  $EX_2$  type.

**TABLE 12.1** Melting and Boiling Points for Group IV Halides

Compound		$EX_2$		$EX_4$	
		m.p. ( $^\circ C$ )	b.p. ( $^\circ C$ )	m.p. ( $^\circ C$ )	b.p. ( $^\circ C$ )
Si	X = F	dec	—	−90.2	−86 subl
	X = Cl	dec	—	−68.8	57.6
	X = Br	—	—	5.4	153
	X = I	—	—	120.5	287.5
Ge	X = F	111	dec	−37 subl	—
	X = Cl	dec	—	−49.5	84
	X = Br	122 dec	—	26.1	186.5
	X = I	dec	—	144	440 dec
Sn	X = F	704 subl	—	705 subl	—
	X = Cl	246.8	652	−33	114.1
	X = Br	215.5	620	31	202
	X = I	320	717	144	364.5
Pb	X = F	855	1290	—	—
	X = Cl	501	950	−15	105 expl
	X = Br	373	916	—	—
	X = I	402	954	—	—

subl = sublimates  
dec = decomposes  
expl = explodes

In general, the ability of the halides to function as Lewis acids results in their being able to react with halide ions (Lewis bases) to form complex ions. Examples of this type of behavior are illustrated by the following equations:



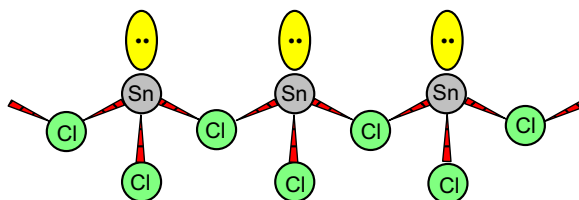
This tendency is greater for the halides of the elements in the +4 oxidation state than it is for those of the +2 state owing to the greater Lewis acid strength of the  $\text{EX}_4$  compounds.

### 12.6.1 The +2 Halides

The dihalides of Si and Ge are polymeric solids that are relatively unimportant compared to those of Sn and Pb. The latter elements are metallic in character and have well-defined +2 oxidation states. Physical data for the divalent halides are shown in Table 12.1. The compounds of Si(II) are relatively unstable because the reaction



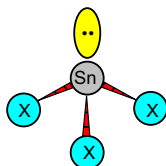
is thermodynamically favored. In the solid state, the  $\text{SnX}_2$  compounds (where X = Cl, Br, or I) exist as chains,



Because of the Lewis acidity of the  $\text{SnX}_2$  compounds, they will react with additional halide ions to form complexes as illustrated by the equation



These complexes have the pyramidal structure shown as



Salts containing these ions can be isolated as solids with large cations such as  $\text{R}_4\text{P}^+$ . Because they have an unshared pair of electrons on the Sn atom, these ions are Lewis bases that form adducts with Lewis acids such as boron halides.



As a result of their being able to react as Lewis acids, compounds having the formula  $\text{SnX}_2$  also form many complexes with molecules containing O and N as the electron donor atoms, e.g.,  $\text{H}_2\text{O}$ , pyridine, amines, and aniline.

Germanium dihalides are obtained by the reaction



The dibromide can be prepared by the following reactions:



However, it disproportionates when heated as represented by the equation



Much of the chemistry of the  $\text{PbX}_2$  and  $\text{SnX}_2$  compounds is that expected for metal halides that are predominantly ionic.

### 12.6.2 The +4 Halides

Not all of the tetrahalides of the Group IVA elements are stable. For example, Pb(IV) is such a strong oxidizing agent that it oxidizes  $\text{Br}^-$  and  $\text{I}^-$  so that  $\text{PbBr}_4$  and  $\text{PbI}_4$  are not stable compounds, and even  $\text{PbCl}_4$  explodes when it is heated. In general, the compounds can be prepared by direct combination of the elements, although the reactions with fluorine may be violent. The other tetrahalides of silicon are produced by passing the halogen over silicon at elevated temperatures.



The behavior of  $\text{SiF}_4$  as a Lewis acid enables it to add two additional fluoride ions to form the hexafluorosilicate ion.



However, the analogous reactions between  $\text{SiCl}_4$ ,  $\text{SiBr}_4$ , and  $\text{SiI}_4$  and the corresponding halide ions do not take place. The reason is primarily a size factor, and the Si is too small to effectively bond to six  $\text{Cl}^-$ ,  $\text{Br}^-$ , or  $\text{I}^-$  ions. However, Ge, Sn, and Pb all form  $\text{EF}_6^{2-}$  and  $\text{ECl}_6^{2-}$  complexes. Because the tetrahalides are Lewis acids, they form complexes with numerous Lewis bases (e.g., amines and phosphines) as illustrated by the reaction



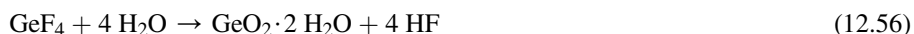
Germanium tetrabromide can be prepared by the reaction of the elements,



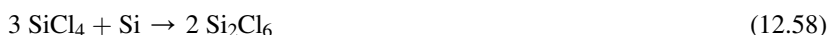
The tetrafluoride is obtained by the reaction



The tetrahalides of the Group IVA elements hydrolyze to give the oxides, hydroxides, or hydrous oxides ( $\text{EO}_2$ ,  $\text{E}(\text{OH})_4$ , or  $\text{EO}_2 \cdot 2 \text{H}_2\text{O}$ ). Typical processes of this type are represented by the equations



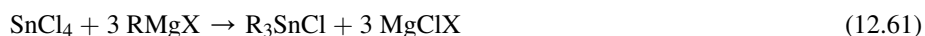
Because Si shows a tendency to form Si–Si bonds, the compounds  $\text{Si}_2\text{X}_6$  are known. For example, the reaction of  $\text{SiCl}_4$  with Si produces  $\text{Si}_2\text{Cl}_6$ .



## 12.7 ORGANIC COMPOUNDS

Although the chemistry of organometallic compounds will be discussed more fully in Chapter 22, the subject will be introduced briefly here. Specific examples of reactions will be given in most cases, but the reactions should be viewed as reaction types that can be carried out using appropriate compounds of Si, Ge, Sn, or Pb. As a result, reactions of one element may be illustrated, but these reactions have been used to develop an extensive organic chemistry of all the elements.

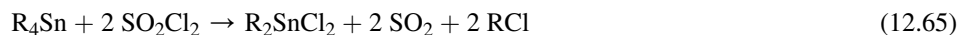
The Group IVA tetrahalides undergo Grignard reactions to produce a large number of compounds containing organic substituents. For example, alkylation reactions such as the following are quite important:



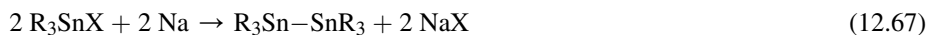
Other alkylating agents such as lithium alkyls ( $\text{LiR}$ ) can be employed.



The mixed alkyl halides can be prepared by the following reactions:



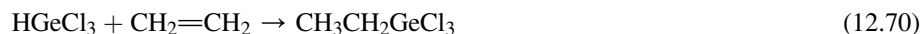
A coupling reaction can be brought about by the reaction of the  $\text{R}_3\text{SnX}$  with sodium.



Silicon reacts with HCl at elevated temperatures to give  $\text{HSiCl}_3$  with the liberation of  $\text{H}_2$ .



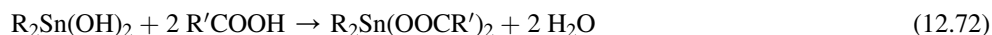
An important reaction of  $\text{HSiCl}_3$  and  $\text{HGeCl}_3$  is their addition across double bonds in alkenes. This reaction is known as the *Speier reaction*, and it leads to alkyl trichloro derivatives.



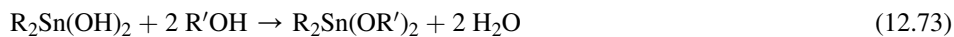
The partially alkylated derivatives can have the remaining halogens removed by hydrolysis to form intermediates that are used in the production of other organic derivatives. A reaction of this type can be shown as follows:



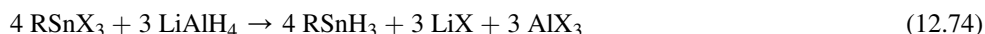
Compounds of this type will react with organic acids,



or with alcohols.



Alkyl hydrides of tin have been prepared by the following reaction:

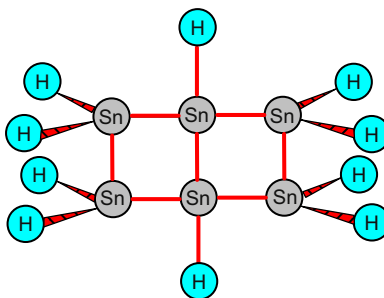


By utilizing combinations of these types of reactions, an extensive organic chemistry of all of the Group IVA elements can be developed.

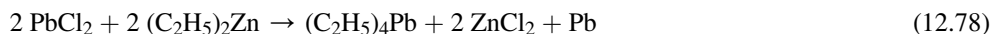
In recent years, a novel type of tin compound has been produced in which the Sn-Sn bonds form rings analogous to those of organic compounds. For example, the triangular ring  $(\text{R}_2\text{Sn})_3$  is produced in the reaction



Other hydride derivatives containing multiple rings are also known. Of the numerous cases of that type, an interesting example is  $\text{Sn}_6\text{H}_{10}$ , and it has the structure



The organic chemistry of lead has had significant economic importance. For example, tetraethyllead was formerly used as an antiknock agent in gasoline to the extent of about 250,000 tons annually. Organic compounds of lead can be prepared by reactions such as the following:



Numerous mixed alkyl and aryl halides of lead are also known, and reactions with water, alcohols, amines, etc., can be used to prepare a large number of other derivatives.

## 12.8 MISCELLANEOUS COMPOUNDS

In this section, a brief overview will be presented of a few other compounds that have not been described in previous sections. Because it can function as a nonmetal, silicon forms silicides with several metals. These materials are often considered as alloys in which the metal and silicon atoms surround each other in a pattern that may lead to unusual stoichiometry. Examples of this type are  $Mo_3Si$  and  $TiSi_2$ . In some silicides, the Si–Si distance is about 235 pm, a distance that is quite close to the value of 234 pm found in the diamond-type structure of elemental silicon. This indicates that the structure contains  $Si_2^{2-}$ , and  $CaSi_2$  is a compound of this type. This compound is analogous to calcium carbide,  $CaC_2$  (actually an acetylide that contains  $C_2^{2-}$  ions (see Chapter 11)).

Silicon combines with carbon to form silicon carbide or *carborundum*,  $SiC$ , and forms are known that have the wurtzite and zinc blende structures. It is a very hard, tough material that is used as an abrasive and a refractory. The powdered material is crushed after mixing with clay and heated in molds to make grinding wheels. It is prepared by the reaction



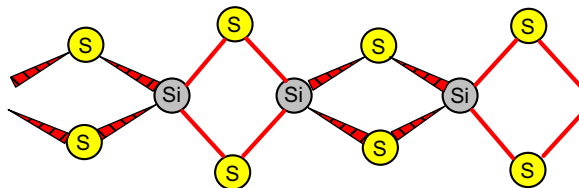
There is also a silicon nitride,  $Si_3N_4$ , which is prepared in an electric furnace by the reaction



Germanium also forms a +2 compound with nitrogen,  $Ge_3N_2$ , as well as an imine,  $GeNH$ . These compounds are obtained by the reactions



Sulfide compounds are formed by all of the Group IVA elements, and lead is found as the sulfide in its principle ore *galena* that has  $Pb^{2+}$  and  $S^{2-}$  ions in a sodium chloride lattice (see Chapter 4). A chain structure is shown by  $SiS_2$  in which each Si is surrounded by four S atoms in an approximately tetrahedral environment.



In addition, sulfide complexes are formed by the reaction





Both GeS and GeS<sub>2</sub> are known and they can be obtained by means of the following reactions:



Hydroxo complexes include those such as Sn(OH)<sub>3</sub><sup>-</sup> that is produced by the following reaction in basic solution:



Solid compounds containing the Sn(OH)<sub>3</sub><sup>-</sup> ion can be obtained. Solids containing the Sn(OH)<sub>6</sub><sup>2-</sup> anion can also be isolated, but they easily lose water to give stannates.

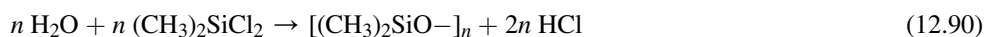


Many of the common compounds such as nitrates, acetates, sulfates, and carbonates are known that contain the elements (particularly Sn and Pb) in the +2 oxidation state, but they will not be discussed further because their behavior is essentially similar that of other compounds containing those ions.

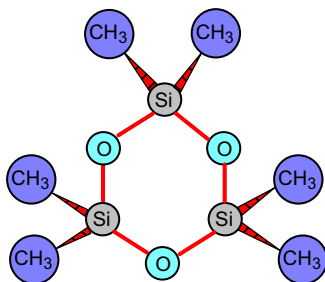
Silicon most often forms single bonds and, as a result, when it bonds to oxygen it bonds to four different atoms. This type of bonding gives rise to the numerous chain and sheet structures that were described earlier. However, the bonds need not all be to oxygen atoms, and many compounds are known in which alkyl groups are bonded to the silicon. The reaction



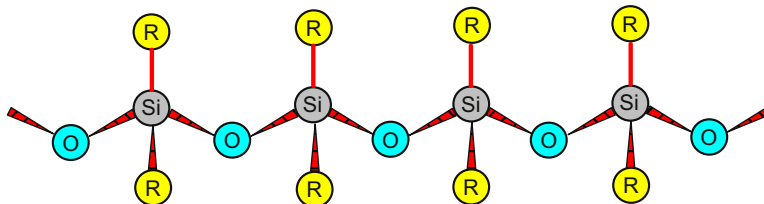
gives a product that contains highly reactive Si–Cl covalent bonds. For example, the hydrolysis reaction of this compound takes place readily.



Silicones are polymeric molecules that have the same silicon–oxygen bridges as are present in the pyroxenes. However, they also have organic groups attached to the silicon atoms rather than bridging oxygen atoms. If  $n = 3$ , as shown in Eq. (12.90), the product is a cyclic trimer that has the structure



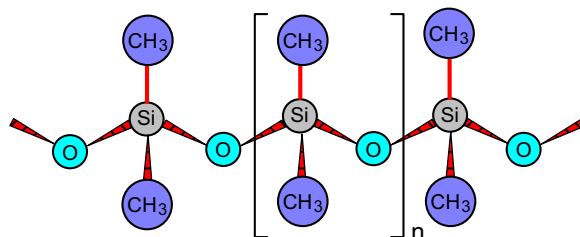
Linear polymers can be obtained by means of the reaction shown in Eq. (12.90) in the presence of dilute sulfuric acid. The linear polymer has a structure that can be shown as



In some cases, there may be fewer than two alkyl groups attached to each silicon atom so that other oxygen bridges form between chains.

The systematic names of the silicones are based on the name *siloxane* because the chain contains silicon atoms, oxygen atoms and has only single bonds. The type of R groups that are present are then indicated using the usual organic

nomenclature. Therefore, the simplest of the silicones is poly(dimethyl)siloxane in which the two R groups attached to silicon atoms are  $\text{CH}_3$  groups.



Silicones are noted for their high thermal stability, water repellency, and lubricating properties. They are used in a wide range of products including personal care items, paints, electronics, and construction materials. As with other polymers, the properties and uses of silicones are dependent on factors such as the types of R groups, length of chain, and extent of cross-linking.

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## PROBLEMS

- Given the bond enthalpies in  $\text{kJ mol}^{-1}$

$$\text{C}-\text{O}, 335; \quad \text{C}=\text{O}, 707; \quad \text{Si}-\text{O}, 464$$

explain why  $\text{CO}_2$  exists as discrete molecules but  $\text{SiO}_2$  does not. Estimate the strength of the  $\text{Si}=\text{O}$  bond.

- Explain why the boiling point of  $\text{GeCl}_4$  is  $84^\circ\text{C}$  but that of  $\text{GeBr}_4$  is  $186.5^\circ\text{C}$ .
- In terms of molecular structure, explain why  $\text{SiCl}_4$  is a Lewis acid, but  $\text{CCl}_4$  is not.
- Using equations, explain why an aqueous solution of  $\text{GeO}_2$  is weakly acidic.
- Write complete, balanced equations for the following processes.
  - Roasting galena in air
  - The preparation of silane from silicon
  - The preparation of germane
  - The preparation of  $\text{Ge}(\text{C}_2\text{H}_5)_4$
  - The preparation of  $(\text{CH}_3)_2\text{SnCl}_2$

6. Explain why the boiling point of  $\text{SnCl}_2$  is  $652^\circ\text{C}$  but that of  $\text{SnCl}_4$  is  $114^\circ\text{C}$ .
7. Draw the structures for the following.
- $(\text{R}_2\text{Sn})_3$
  - $\text{SnS}_2$
  - $\text{Si}_2\text{O}_7^{6-}$
8. Write complete, balanced equations for the processes indicated.
- The combustion of silane
  - The preparation of  $\text{PbO}_2$
  - The reaction of  $\text{CaO}$  with  $\text{SiO}_2$  at high temperature
  - The preparation of  $\text{GeCl}_2$
  - The reaction of  $\text{SnCl}_4$  with water
  - The preparation of tetraethyllead
9. Why is  $\text{SnCl}_3^-$  a Lewis base but  $\text{SnCl}_4$  behaves as a Lewis acid?
10. (a) Write the equation to show the reaction of  $(\text{CH}_3)_3\text{SiCl}$  with water.  
(b) Write an equation for the reaction that occurs when the product in part (a) is heated.
11. Complete and balance the following.
- $\text{Ge} + \text{HNO}_3 \rightarrow$
  - $\text{SnO}_2 + \text{CaO} \xrightarrow{\text{heat}}$
  - $\text{SnCl}_4 + \text{HCl} \rightarrow$
  - $\text{HSiCl}_3 + \text{CH}_3\text{CH}=\text{CHCH}_3 \rightarrow$
  - $(\text{C}_2\text{H}_5)_2\text{SnCl}_2 + \text{H}_2\text{O} \rightarrow$
12. Complete and balance the following.
- $\text{SnBr}_4 + \text{LiC}_4\text{H}_9 \rightarrow$
  - $\text{GeH}_4 + \text{O}_2 \rightarrow$
  - $\text{SnO} + \text{NaOH} \rightarrow$
  - $\text{SiCl}_2 \xrightarrow{\text{heat}}$
  - $\text{GeCl}_4 + \text{P}(\text{C}_2\text{H}_5)_3 \rightarrow$
13. In the  $\text{Si}_3\text{O}_9^{6-}$  ion, determine the bond character of each Si—O bond (see Chapter 4).
14. Describe the preparation, uses, and structures of silicone polymers.
15. Describe two distinctly different compounds that have the formula  $\text{PbO}_2$ .

# Nitrogen

Nitrogen is an abundant element (78% by volume of the atmosphere) that was discovered in 1772 by several workers independently. There are also many important naturally occurring compounds that contain nitrogen, especially the nitrates  $\text{KNO}_3$  and  $\text{NaNO}_3$ . Of course, the fact that all living protoplasm contains amino acids means that it is also an essential element for life. The range and scope of nitrogen chemistry is enormous, so it is one of the most interesting elements.

## 13.1 ELEMENTAL NITROGEN

The nitrogen atom has a valence shell population of  $2s^2 2p^3$  so it has a  $^4S$  ground state. Elemental nitrogen consists of diatomic molecules so the element is frequently referred to as dinitrogen. The atoms are found to consist of two isotopes,  $^{14}\text{N}$  (99.635%) and  $^{15}\text{N}$ . This is somewhat unusual in that  $^{14}\text{N}$  is an odd–odd nuclide containing seven protons and seven neutrons, and there are very few stable odd–odd nuclei.

Nitrogen molecules are small with a bond length of only  $1.10 \text{ \AA}$  (110 pm) owing to the strong triple bond holding the atoms together. The force constant is correspondingly large, being  $22.4 \text{ mdyne \AA}^{-1}$ . The molecular orbital diagram for  $\text{N}_2$ , shown in Figure 13.1, indicates that the bond order is 3 in this extremely stable molecule.

Nitrogen is a rather unreactive element, and the reason is that the  $\text{N}\equiv\text{N}$  bond energy is  $946 \text{ kJ mol}^{-1}$ . This lack of reactivity is somewhat unlike other nonmetals given the position of the atom in the periodic table and the fact that nitrogen is a nonmetal having an electronegativity of 3.0 (the third highest value). As a result of the stability of the  $\text{N}_2$  molecule, many nitrogen compounds are unstable, some explosively so.

Elemental nitrogen is commercially obtained by the distillation of liquid air because oxygen boils at  $-183 \text{ }^\circ\text{C}$  and liquid nitrogen boils at  $-195.8 \text{ }^\circ\text{C}$ . Because of the large amount of oxygen required (especially in processes for making steel), this constitutes the major industrial source of nitrogen. Small amounts of nitrogen can be obtained in the laboratory by the decomposition of sodium azide,  $\text{NaN}_3$ .



A different decomposition reaction can be shown as follows:

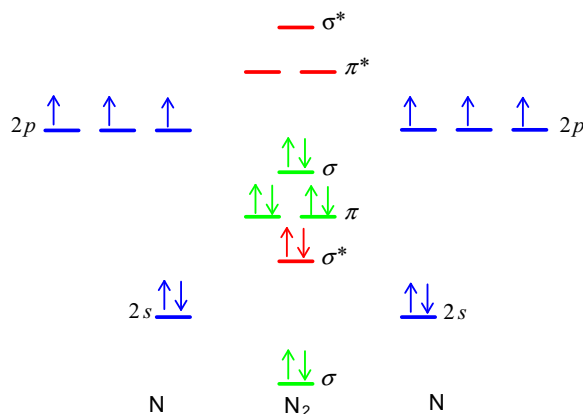
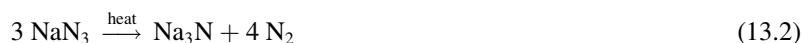


FIGURE 13.1 Molecular orbital diagram for  $\text{N}_2$ .

The element can also be obtained by the careful decomposition of ammonium nitrite,



It is also possible to burn phosphorus in a closed container to remove oxygen and leave nitrogen, or to remove oxygen from air by reaction with pyrogallol leaving nitrogen. However, nitrogen prepared in these ways is not of high purity.

## 13.2 NITRIDES

Although nitrogen is somewhat unreactive, at high temperatures it reacts with metallic elements to produce some binary compounds. For example, the reaction with magnesium can be shown as



As for hydrides, borides, and carbides, different types of nitrides are possible depending on the type of metallic element. The classifications of nitrides are similarly referred to as ionic (saltlike), covalent, and interstitial. However, it should be noted that there is a transition of bond types. Within the covalent classification, nitrides are known that have a diamond or graphite structure. Principally, these are the boron nitrides that were discussed in Chapter 9.

Nitrogen forms binary compounds with most other elements, although many of them are not obtained by direct combination reactions. The ionic compounds, e.g.,  $\text{Mg}_3\text{N}_2$  or  $\text{Na}_3\text{N}$ , result when the difference between the electronegativities of the two atoms is about 1.6 units or greater. These compounds are usually prepared by direct combination of the elements as illustrated in Eq. (13.4). The reaction with sodium can be represented as



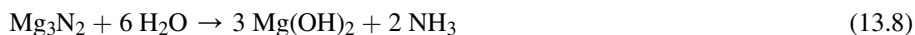
although the reaction is not actually this simple. However, just as oxides result from the decomposition of hydroxides,



nitrides result from the decomposition of amides.



Compounds containing the  $\text{N}^{3-}$  ion readily react with water owing to the strongly basic nature of this ion.



Most nonmetals produce one or more covalent “nitrides.” This name implies that the nitrogen atom is the more electronegative one, so compounds such as  $\text{NF}_3$  or  $\text{NO}_2$  are excluded. However, nitrogen has a higher electronegativity than most other nonmetals so there are many covalent nitrides and they have enormously varied properties. For example compounds of this type include such compounds as  $\text{HN}_3$ ,  $\text{NH}_3$ ,  $\text{S}_4\text{N}_4$ ,  $(\text{CN})_2$ , and numerous others. Because of the greatly differing character of these compounds, no general methods of preparation can be given. Accordingly, some of the preparations will be given in the discussion of nitrogen compounds with specific elements.

Many transition metals form interstitial nitrides by reaction of the metal with  $\text{N}_2$  or  $\text{NH}_3$  at elevated temperatures. Nitrogen atoms occupy some fraction of the interstitial positions in these metals so that the “compounds” frequently deviate from exact stoichiometry. Rather, a range of materials exists with the composition obtained depending on the temperature and pressure used in the synthesis. Predictable changes in physical properties result from placing nitrogen atoms in interstitial positions in metals. These metal nitrides are hard, brittle, high-melting solid materials that generally have a metallic appearance. As in the case of interstitial hydrides, the composition of these materials may approach a simple ratio of metal to nitrogen, but the actual composition depends on the experimental conditions.

## 13.3 AMMONIA AND AQUO COMPOUNDS

One of the interesting features of the chemistry of nitrogen compounds is the parallel that exists between the chemistry of ammonia and water. For example, both molecules can be protonated, and the products are  $\text{NH}_4^+$  and  $\text{H}_3\text{O}^+$ . Both can also be deprotonated, and the products are  $\text{NH}_2^-$  and  $\text{OH}^-$ . Solid compounds can be prepared that contain these ions. Also, both water and ammonia can be completely deprotonated giving rise to the ions  $\text{O}^{2-}$  and  $\text{N}^{3-}$ , both of which are

**TABLE 13.1** The Ammono and Aquo Series of Compounds

Ammono Species	Aquo Species
NH <sub>4</sub> <sup>+</sup>	H <sub>3</sub> O <sup>+</sup>
NH <sub>3</sub>	H <sub>2</sub> O
NH <sub>2</sub> <sup>-</sup>	OH <sup>-</sup>
NH <sup>2-</sup>	O <sup>2-</sup>
N <sup>3-</sup>	—
H <sub>2</sub> N—NH <sub>2</sub>	HO—OH
RNH <sub>2</sub>	ROH
RNHR	ROR
R <sub>3</sub> N	—
HN=NH	—
NH <sub>2</sub> OH	—

exceedingly strong bases. Many of the derivatives of ammonia and water are summarized in Table 13.1. Discussions of the similarities will appear in subsequent sections.

## 13.4 HYDROGEN COMPOUNDS

### 13.4.1 Ammonia

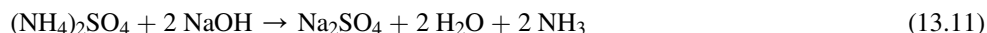
The most common and important nitrogen–hydrogen compound is ammonia. Because liquid ammonia is a commonly used nonaqueous solvent, it was discussed in Section 6.2.3, and its properties are listed in Table 6.5. Approximately 22 billion pounds of NH<sub>3</sub> are used annually, mostly as fertilizer or as the starting material for preparing nitric acid. The *Haber process* is used for the synthesis of NH<sub>3</sub> from the elements.



The conditions used represent a compromise because the reaction is faster at higher temperatures, but NH<sub>3</sub> also becomes less stable as a result of its heat of formation being  $-46 \text{ kJ mol}^{-1}$ . One of the catalysts used is alpha iron containing some iron oxide and other oxides (MgO and SiO<sub>2</sub>) that increases the active surface by expansion of the lattice. Considerable amounts of ammonia are also obtained during the heating of coal to produce coke because this results in the decomposition of organic nitrogen compounds. The hydrolysis of calcium cyanamide, CaCN<sub>2</sub>, also produces NH<sub>3</sub>.



Small quantities of NH<sub>3</sub> can be prepared in the laboratory by the reaction of ammonium salts with strong bases.



The reaction of an ionic nitride, e.g., Mg<sub>3</sub>N<sub>2</sub>, with water can also be used.

Ammonia is a colorless gas with a characteristic odor. The compound freezes at  $-77.8 \text{ }^\circ\text{C}$  and boils at  $-33.35 \text{ }^\circ\text{C}$ . Hydrogen bonding is extensive in the liquid and solid states because of the polarity of the N–H bonds and the unshared pair of electrons on the nitrogen atom. The structure of the NH<sub>3</sub> molecule has been described in Chapter 3. Because it has an unshared pair of electrons on the nitrogen atom, it is both a Brønsted and Lewis base that forms many complexes with metals.

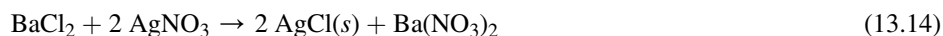
Ammonia is a weak base, having  $K_b = 1.8 \times 10^{-5}$ , and it ionizes in water according to the following equation:



The formula  $\text{NH}_4\text{OH}$  is sometimes used for convenience, but such a “molecule” apparently does not exist or at least is unstable. Ammonia is extremely soluble in water, however. Most of the  $\text{NH}_3$  is physically dissolved in water with a small amount undergoing the reaction shown in Eq. (13.12). Ammonia can also act as an acid toward extremely strong Brønsted bases.



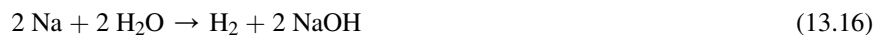
Liquid ammonia has been extensively used as a nonaqueous solvent (see Chapter 6), but only a few aspects of that chemistry will be mentioned here. Many inorganic salts are appreciably soluble in the liquid owing to its polarity although reactions are frequently different than they are in water. For example, in water,



solid  $\text{AgCl}$  will form because  $\text{AgCl}$  is insoluble in water. However, in liquid  $\text{NH}_3$ , the reaction is



because of the insolubility of  $\text{BaCl}_2$ . Thus,  $\text{AgCl}$  is soluble in liquid ammonia as it also is in aqueous ammonia solutions. The numerous similarities between the aquo and ammono series of compounds have already illustrated (see Table 13.1). One great difference involves the reactivity of Group IA metals in liquid  $\text{NH}_3$  versus their reactivity in  $\text{H}_2\text{O}$ . Group IA metals react vigorously and rapidly with water



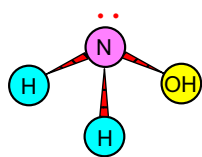
However, the similar reactions with ammonia



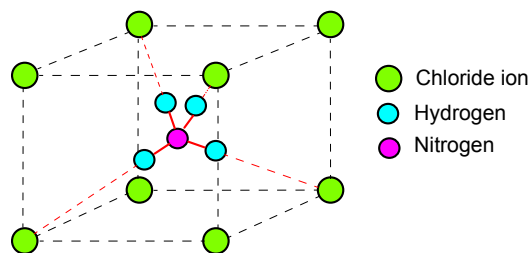
take place very slowly. The amide ion is a stronger base than  $\text{OH}^-$  and some reactions requiring a strongly basic medium take place more readily in liquid ammonia than they do in water. Refer to Chapter 6 for other aspects of the chemistry of liquid  $\text{NH}_3$ .

Ammonium salts frequently resemble those of potassium, rubidium, or cesium in morphology and solubility because of the similar sizes of the ions. The radii of these species are as follows:  $\text{K}^+ = 133 \text{ pm}$ ,  $\text{Rb}^+ = 148 \text{ pm}$ ,  $\text{Cs}^+ = 169 \text{ pm}$ , and  $\text{NH}_4^+ = 148 \text{ pm}$ . For example, the unit cell of ammonium chloride has the structure shown in Figure 13.2.

A derivative of ammonia that is useful in numerous ways is hydroxylamine,  $\text{NH}_2\text{OH}$ . The compound has a melting point of  $33^\circ\text{C}$  and may decompose explosively when heated. The structure of the molecule can be shown as



As suggested by the presence of an unshared pair of electrons on the nitrogen atom, hydroxylamine is a weak base that has a  $pK_b$  of 3.37 whereas that of ammonia is 4.75. As with other amine bases, hydroxylamine readily forms acid salts such as



**FIGURE 13.2** The structure of ammonium chloride. Note the hydrogen bonds between the hydrogen atoms in the ammonium ion and the chloride ions.

hydroxylamine hydrochloride,  $\text{HONH}_3^+\text{Cl}^-$ , a form that is commercially available. Hydroxylamine behaves as an effective reducing agent in many reactions involving the preparation of polymers, medicines, and dyestuffs. It can be prepared by several routes including the hydrolysis of nitroalkanes in the presence of a strong acid.



Hydroxyurea is produced by the reaction of hydroxylamine with sodium cyanate.



The medicinal uses of hydroxyurea will be described more fully in Chapter 23.

Ammonium nitrate can be decomposed *carefully* to produce  $\text{N}_2\text{O}$ ,



and ammonium nitrite can be decomposed to produce  $\text{N}_2$ .



Ammonium nitrate will explode violently if the reaction is initiated by another primary explosive. Mixtures of  $\text{NH}_4\text{NO}_3$  and TNT (2,4,6-trinitrotoluene) are known as *amatol*, a military explosive. Most ammonium salts can be decomposed by heating, but many solid compounds that contain the ammonium cation and an anion that is the conjugate of a weak acid decompose quite readily with only mild heating. Some examples are illustrated in the following equations:



The most important reaction of  $\text{NH}_3$  is its oxidation by the *Ostwald process*.



The  $\text{NO}$ , a reactive gas, can easily be oxidized,



Nitric acid is produced when  $\text{NO}_2$  disproportionates in water.



The nitric acid can then be concentrated by distillation to about 68% by weight (to give a solution called “concentrated” nitric acid). This is the source of almost all nitric acid. The  $\text{HNO}_2$  produced is unstable,



and the gases are recycled (see Section 13.7.3).

### 13.4.2 Hydrazine, $\text{N}_2\text{H}_4$

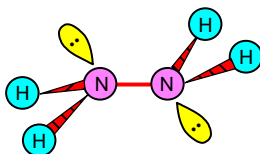
This compound actually has the structure  $\text{H}_2\text{N}-\text{NH}_2$ , and it is sort of an analogue of hydrogen peroxide,  $\text{HO}-\text{OH}$ . Both compounds are thermodynamically unstable, the heat of formation of hydrazine being  $+50\text{ kJ mol}^{-1}$ . Hydrazine is a weak diprotic base having  $K_{b1} = 8.5 \times 10^{-7}$  and  $K_{b2} = 8.9 \times 10^{-16}$ . It is a good reducing agent and reacts vigorously with strong oxidizing agents and readily burns. For example, the reaction



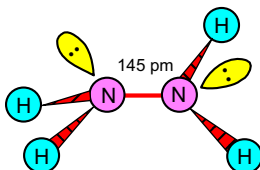
is strongly exothermic. A substituted hydrazine,  $(\text{CH}_3)_2\text{N}-\text{NH}_2$ , has been used as a rocket fuel as has hydrazine itself.



It is interesting to note that the molecule  $\text{N}_2\text{H}_4$  is polar ( $\mu = 1.75 \text{ D}$ ) so the structure is not correctly represented as having a *trans* orientation shown as

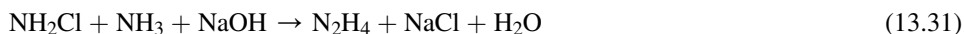


Instead, the structure of the molecule is represented as



with an N–N bond length of 145 pm. The protonated species,  $\text{N}_2\text{H}_5^+$  actually has a shorter N–N bond length because attaching  $\text{H}^+$  to one of the unshared pairs of electrons reduces the repulsion between it and the other unshared pair.

The primary synthesis of hydrazine is the *Raschig process*, and the first step in the process leads to chloramine,  $\text{NH}_2\text{Cl}$ .



However, there is also a competing reaction that takes place and it can be represented as



The gelatin binds to traces of metal ions that catalyze this reaction, and it also catalyzes the reaction shown in Eq. (13.30).

### 13.4.3 Diimine, $\text{N}_2\text{H}_2$

This compound is actually  $\text{HN}=\text{NH}$ . It is not stable but it appears to exist at least as an intermediate in some processes. It decomposes to give  $\text{N}_2$  and  $\text{H}_2$ .



Diimine is prepared by the reaction of chloramine with a base.

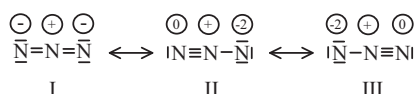


### 13.4.4 Hydrogen Azide, $\text{HN}_3$

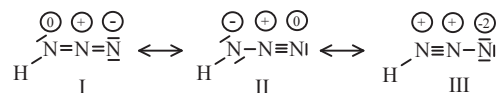
Hydrogen azide (or hydrazoic acid) is a volatile compound (m.p.  $-80^\circ\text{C}$ , b.p.  $37^\circ\text{C}$ ) that is a weak acid having  $K_a = 1.8 \times 10^{-5}$ . It is a dangerous explosive (it contains 98% nitrogen!), and it is highly toxic. Other covalent azides such as  $\text{CH}_3\text{N}_3$  and  $\text{ClN}_3$  are also explosive. Heavy metal salts such as  $\text{Pb}(\text{N}_3)_2$  and  $\text{AgN}_3$  are also sensitive to shock and have been used as primary explosives (detonators). In contrast, ionic azides such as  $\text{Mg}_2\text{N}_3$  and  $\text{NaN}_3$  are relatively stable and decompose slowly upon heating strongly.



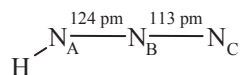
The marked difference in stability of ionic and covalent azides is sometimes explained in terms of their structural differences. For example, the azide ion,  $\text{N}_3^-$  is a linear triatomic species that has 16 valence electrons, and it has three contributing resonance structures that can be shown as follows:



All of these structures contribute to the true structure although structure I is certainly the dominant one. However, a covalent azide such as  $\text{HN}_3$  represents a somewhat different situation as shown by the following structures.

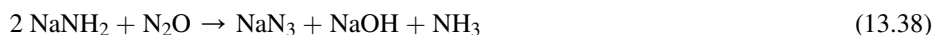


Structure III is highly unfavorable because of the identical formal charges on adjacent atoms and the higher formal charges. A consideration of the formal charges in structures I and II would lead one to predict that the two structures contribute about equally. In  $\text{HN}_3$ , the bond lengths are



Bonds between nitrogen atoms have typical lengths as follows:  $\text{N}-\text{N}$ , 145 pm;  $\text{N}=\text{N}$ , 125 pm;  $\text{N}\equiv\text{N}$ , 110 pm. Because structures I and II place single and double bonds between  $\text{N}_A$  and  $\text{N}_B$ , the observed bond length would be expected to be somewhere between that expected for a single bond and that expected for a double bond (125 and 145 pm). It is, in fact, about the same as a double bond, which means that structures I and II do not contribute quite equally to the actual structure. Examination of the bond distance between atoms  $\text{N}_B$  and  $\text{N}_C$  (close to  $\text{N}\equiv\text{N}$ ) leads to a similar conclusion. Thus, although there are three resonance structures that contribute to the structure of the azide ion, only two structures contribute to the structure of a covalent azide. It is generally true that the greater the number of contributing resonance structures, the more stable the species (the lower the energy of the species).

Sodium azide can be prepared as follows:



Aqueous solutions of  $\text{HN}_3$  can be obtained by acidifying a solution of  $\text{NaN}_3$  because the acid is only slightly dissociated. A dilute aqueous solution of  $\text{HN}_3$  can also be prepared by the reaction



The azide ion is a good ligand, and it forms numerous complexes with metal ions. Chlorazide ( $\text{ClN}_3$ ) is an explosive compound prepared by the reaction of  $\text{OCl}^-$  and  $\text{N}_3^-$ . As in the case of  $\text{CN}^-$ , the azide ion is a pseudohalide ion. Pseudohalogens are characterized by the formation of an insoluble silver salt, the acid  $\text{H}-\text{X}$  exists,  $\text{X}-\text{X}$  is volatile, and they combine with other pseudohalogens to give  $\text{X}-\text{X}'$ . Although such pseudohalogens as  $(\text{CN})_2$  exist because of the oxidation of the  $\text{CN}^-$  ion,



the corresponding  $\text{N}_3-\text{N}_3$  is unknown. It would, after all, be an allotrope of nitrogen and the  $\text{N}_2$  form is naturally much more stable.

## 13.5 NITROGEN HALIDES

### 13.5.1 $\text{NX}_3$ Compounds

The nitrogen halides have the general formula  $\text{NX}_3$ , but not  $\text{NX}_5$  as in the case of phosphorus. However, some mixed halides such as  $\text{NF}_2\text{Cl}$  are known. Fluorine also forms  $\text{N}_2\text{F}_4$  and  $\text{N}_2\text{F}_2$  (these compounds will be discussed later) that are analogous to hydrazine and diimine, respectively. Except for  $\text{NF}_3$  (b.p.  $-129^\circ \text{C}$ ), the compounds are explosive. Because of this behavior, most of what is presented here will deal with the fluorine compounds.

For  $\text{NF}_3$ ,  $\Delta H_f^\circ = -109 \text{ kJ mol}^{-1}$ . The compound does not hydrolyze in water as do most other nonmetal halides, including  $\text{NCl}_3$ .

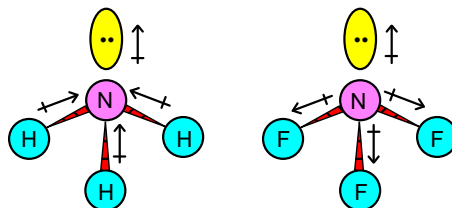


The trichloride compound behaves as though nitrogen is the negative atom. Almost all covalent nonmetal halides (e.g.,  $\text{PCl}_5$ ,  $\text{PI}_3$ ,  $\text{SbCl}_3$ , etc.) react readily with water. One could argue that with about 1.0 units difference between the electronegativities of N and F that this is not a “typical” covalent halide. However,  $\text{SbF}_3$  having an even greater difference in electronegativity between the two atoms (greater ionic character) does hydrolyze. Unlike the chloride, there is no

explosive character to  $\text{NF}_3$ . For  $\text{NCl}_3$ ,  $\Delta H_f^\circ = +232 \text{ kJ mol}^{-1}$ , and its behavior is quite different from that of  $\text{NF}_3$  reflecting this fact.

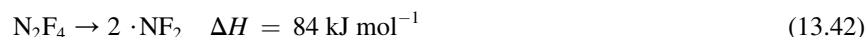
The electrolysis of  $\text{NH}_4\text{F} \cdot \text{HF}$ , which can also be described as  $\text{NH}_4^+ \text{HF}_2^-$ , yields  $\text{NF}_3$ . Also, the reaction of  $\text{NH}_3$  and  $\text{F}_2$  produces  $\text{NF}_3$ ,  $\text{N}_2\text{F}_4$ ,  $\text{N}_2\text{F}_2$ ,  $\text{NHF}_2$ , etc., with the actual distribution of products depending on the reaction conditions.

The dipole moment of  $\text{NF}_3$  is only 0.24 D whereas that of  $\text{NH}_3$  is 1.47 D. The reason for this large difference is that the bond dipoles (represented as arrows with the heads representing the negative ends) are in opposite directions in the two molecules.



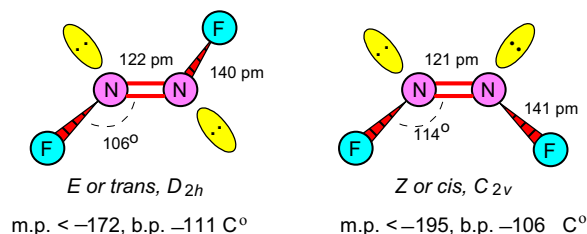
Also, the effect of the unshared pair reinforces the bond moments in  $\text{NH}_3$ , but partially cancels them in  $\text{NF}_3$ . However, the bond angles in  $\text{NH}_3$  and  $\text{NF}_3$  are  $103.3^\circ$  and  $103.2^\circ$ , respectively, so the difference in polarity is due to difference in bond polarities rather than molecular geometry. As a result of these differences in availability of the unshared pairs of electrons,  $\text{NF}_3$  does not act as an electron pair donor.

The reaction of  $\text{NF}_3$  with copper can be used to prepare  $\text{N}_2\text{F}_4$  (b.p.  $-73^\circ \text{C}$ ). In the gas phase,  $\text{N}_2\text{F}_4$  dissociates into  $\cdot\text{NF}_2$ .



### 13.5.2 Difluorodiazine, $\text{N}_2\text{F}_2$

The compound  $\text{N}_2\text{F}_2$ , also known as dinitrogen difluoride, exists in two forms.



The *Z* (*cis*) form is the more reactive of the two. For example, the *Z* form will slowly attack glass to produce  $\text{SiF}_4$  whereas the *E* (*trans*) form will not. The *Z* form is also a good fluorinating agent that reacts with many materials as illustrated by the following equations.



A cation that can be described as  $\text{N}_2\text{F}^+$  can be generated by fluoride ion removal by a strong Lewis acid such as  $\text{SbF}_5$ .



It is also possible to prepare compounds having the general formula  $\text{NH}_n\text{X}_{3-n}$ , but only  $\text{ClNH}_2$ , chloramine, is important.

### 13.5.3 Oxyhalides

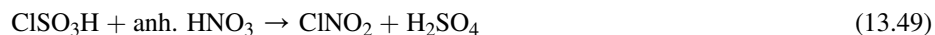
A number of oxyhalides of nitrogen are known including the types  $\text{XNO}$  (*nitrosyl* halides) with  $\text{X} = \text{F}$ ,  $\text{Cl}$ , or  $\text{Br}$ , and  $\text{XNO}_2$  (*nitryl* halides) with  $\text{X} = \text{F}$  or  $\text{Cl}$ . All of these compounds are gases at room temperature. Nitrosyl halides can be prepared by the reactions of halogens with  $\text{NO}$ .



The oxyhalides are very reactive and generally function to halogenate other species. For example, they react to produce halo complexes as illustrated by the equations



Nitryl chloride can be prepared by the reaction



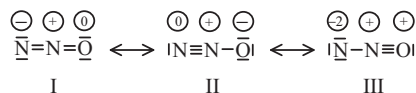
These interesting compounds also undergo halogenation reactions.

## 13.6 NITROGEN OXIDES

There are several oxides of nitrogen that have been well characterized. These are described in [Table 13.2](#).

### 13.6.1 Nitrous Oxide, $\text{N}_2\text{O}$

The  $\text{N}_2\text{O}$  molecule is linear as a result of it being a 16-electron triatomic molecule. Three resonance structures can be drawn for this molecule.



Because three atoms would require 24 electrons for complete octets, there must be eight electrons shared (four pairs, four bonds). Therefore, the central atom will have only the four shared pairs of electrons surrounding it. That precludes oxygen from being in the middle because it would thereby acquire a +2 formal charge. Accordingly, nitrogen is the central atom. The contribution from structure III above would be insignificant because of the high negative formal charge on N and a positive formal charge on O. Structures I and II contribute about equally to the actual structure as is indicated by the fact that the dipole moment of  $\text{N}_2\text{O}$  is only about 0.166 D. Structures I and II would tend to cancel dipole effects if they contributed exactly equally. Bond distances are also useful in this case. The N–N distance is 113 pm and the N–O distance is 119 pm. The bond length in  $\text{N}_2$  is 110 pm (where there is a triple bond) and the N=N bond length is about 125 pm. Therefore, it is apparent that the N–N bond in  $\text{N}_2\text{O}$  is between a double and triple bond, as expected from the contributions of structures I and II.

Nitrous oxide is rather unreactive, but it can function as an oxidizing agent, and it reacts explosively with  $\text{H}_2$ ,



As it does in air, magnesium will burn in an atmosphere of  $\text{N}_2\text{O}$ .



**TABLE 13.2** Oxides of Nitrogen

Formula	Name	Characteristics
$\text{N}_2\text{O}$	Nitrous oxide	Colorless gas, weak oxidizing agent
NO	Nitric oxide	Colorless gas, paramagnetic
$\text{N}_2\text{O}_3$	Dinitrogen trioxide	Blue solid, dissociates in gas
$\text{NO}_2$	Nitrogen dioxide	Brown gas, equilibrium mixture
$\text{N}_2\text{O}_4$	Dinitrogen tetroxide	Colorless gas
$\text{N}_2\text{O}_5$	Dinitrogen pentoxide	Solid is $\text{NO}_2^+\text{NO}_3^-$ , gas unstable

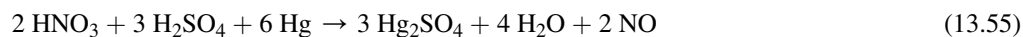
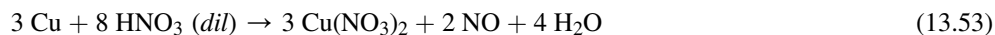
Nitrous oxide is quite soluble in water. At 0 °C, a volume of water dissolves 1.3 times its volume of N<sub>2</sub>O at 1 atm pressure. It is used as a propellant gas in canned whipped cream, and it has been used as an anesthetic (laughing gas). The melting point of N<sub>2</sub>O is -91 °C and the boiling point is -88 °C.

### 13.6.2 Nitric Oxide, NO

Nitric oxide, NO, is an important compound because it is a precursor of nitric acid. It is prepared commercially by the Ostwald process.



In the laboratory, several reactions can be used to produce NO.



The NO molecule has an odd number of electrons as can be seen from the molecular orbital diagram that is shown in Figure 13.3. The molecular orbital diagram shows that the bond order is 2.5 in the NO molecule (bond length 115 pm). If an electron is removed, it comes from the  $\pi^*$  orbital, leaving NO<sup>+</sup> for which the bond order is 3 (bond length 106 pm). The ionization potential for NO is 9.2 eV (888 kJ mol<sup>-1</sup>) so it loses an electron rather easily to very strong oxidizing agents to generate NO<sup>+</sup>, a species that is isoelectronic with N<sub>2</sub>, CN<sup>-</sup>, and CO. The nitrosyl ion is a good coordinating agent and many complexes containing NO<sup>+</sup> as a ligand are known. The NO molecule behaves as a donor of three electrons (one transferred to the metal followed by donation of an unshared pair on the nitrogen atom). Halogens react with NO to produce XNO, nitrosyl halides.



and it is easily oxidized to NO<sub>2</sub>.



This reaction is one of the steps in converting NH<sub>3</sub> to HNO<sub>3</sub>.

The fact that gaseous NO does not dimerize has been the subject of considerable conjecture. If one examines the molecular orbital diagram for NO, it is apparent that the bond order is 2.5. If dimers were formed, a structure shown as

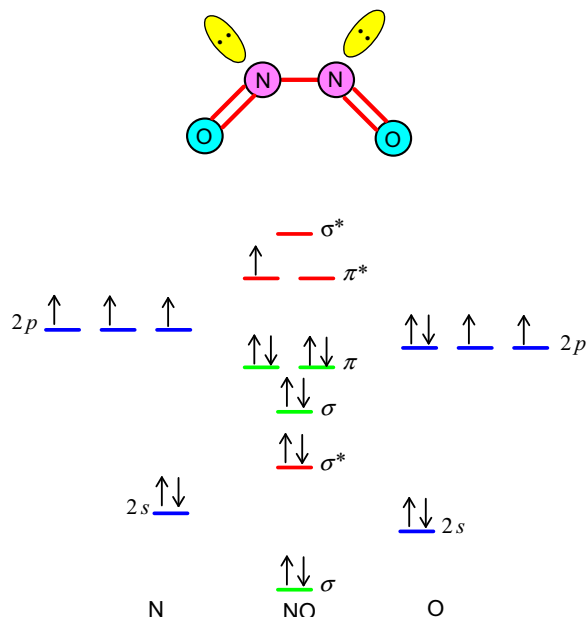


FIGURE 13.3 Molecular orbital diagram for NO.

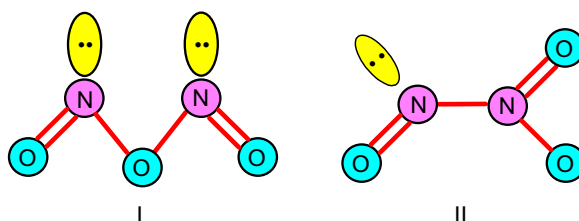
would almost certainly be the dominant resonance structure. This structure contains a total of five bonds corresponding to an average of 2.5 bonds per NO unit. Therefore, there is no net increase in the number of bonds in the dimer over the two separate NO units, and from the standpoint of energy there is little reason for dimers to form. The melting point of NO is  $-164\text{ }^{\circ}\text{C}$  and the boiling point is  $-152\text{ }^{\circ}\text{C}$ . The low boiling point and small liquid range, only about  $12\text{ }^{\circ}\text{C}$ , are indicative of very weak intermolecular forces and suggest little tendency for NO to dimerize. In the solid state, NO does exist as dimers, and the liquid may have some association, but unlike  $\text{NO}_2$ , NO does not dimerize in the gas phase.

### 13.6.3 Dinitrogen Trioxide, $\text{N}_2\text{O}_3$

In some ways,  $\text{N}_2\text{O}_3$  (m.p.  $-101\text{ }^{\circ}\text{C}$ ) behaves as a 1:1 mixture of NO and  $\text{NO}_2$ . In the gas phase, it is largely dissociated,



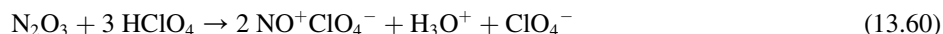
It is prepared by cooling a mixture of NO and  $\text{NO}_2$  at  $-20$  to  $-30\text{ }^{\circ}\text{C}$  or by the reaction of NO with  $\text{N}_2\text{O}_4$ . Two forms of  $\text{N}_2\text{O}_3$  are possible having the structures



The dominant form appears to be II in which the N–N bond length is 186 pm. It reacts with water to produce a solution containing nitrous acid,  $\text{HNO}_2$ .



It also reacts with strong acids such as  $\text{HClO}_4$  and  $\text{H}_2\text{SO}_4$  to produce species containing the  $\text{NO}^+$  ion.

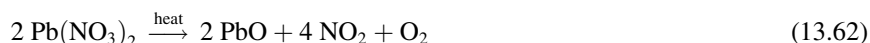


### 13.6.4 Nitrogen Dioxide, $\text{NO}_2$ and $\text{N}_2\text{O}_4$

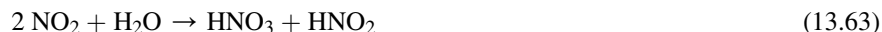
$\text{NO}_2$ , a toxic gas, can be prepared by the oxidation of NO



or the decomposition of  $\text{Pb}(\text{NO}_3)_2$ .



The major importance of the compound is that it reacts with water as one step in the preparation of nitric acid.

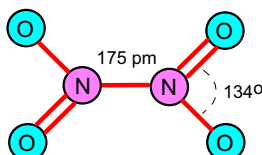


The molecule is angular ( $C_{2v}$ ) with a bond angle of  $134^{\circ}$ .

Unlike NO,  $\text{NO}_2$  extensively dimerizes.

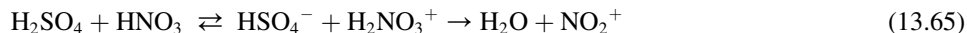


At  $135\text{ }^{\circ}\text{C}$ , the dissociation is 99% but at  $25\text{ }^{\circ}\text{C}$  it is only about 20% dissociated, and in the liquid state (b.p.  $-11\text{ }^{\circ}\text{C}$ ) it is completely associated as  $\text{N}_2\text{O}_4$ . The structure of  $\text{N}_2\text{O}_4$  can be shown as



but other forms such as  $\text{NO}^+\text{NO}_3^-$ ,  $\text{ONONO}_2$ , etc., are indicated under some conditions. The structure shown has a very long N–N bond ( $\sim 175$  pm) whereas in  $\text{N}_2\text{H}_4$  the N–N bond is 147 pm.

The nitronium ion,  $\text{NO}_2^+$ , derived from  $\text{NO}_2$ , is of considerable interest because it is the attacking species in nitration reactions (see Chapter 6). It is generated by the interaction of concentrated sulfuric and nitric acids,



Also, liquid  $\text{N}_2\text{O}_4$  has been extensively studied as a nonaqueous solvent. Autoionization, to the extent that it occurs, appears to be



and in many reactions  $\text{N}_2\text{O}_4$  reacts as if these ions were present. It does not react as if the alternate ionization mode to  $\text{NO}_2^+\text{NO}_2^-$  occurs. Accordingly, compounds such as  $\text{NOCl}$  (actually  $\text{ONCl}$  in structure) are acids and nitrates are bases in liquid  $\text{N}_2\text{O}_4$ . Metals react with  $\text{N}_2\text{O}_4$  to give the nitrates,



where  $\text{M} = \text{Na}, \text{K}, \text{Zn}, \text{Ag},$  or  $\text{Hg}$ . Liquid  $\text{N}_2\text{O}_4$  has a dielectric constant of 2.42 at  $15^\circ\text{C}$ , and this value is comparable to that of many organic compounds. Liquid  $\text{N}_2\text{O}_4$  is an oxidizing agent that has been used in conjunction with a dimethyl hydrazine fuel.

### 13.6.5 Dinitrogen Pentoxide, $\text{N}_2\text{O}_5$

This oxide is the anhydride of nitric acid from which it can be prepared by dehydration using a strong dehydrating agent at low temperatures.



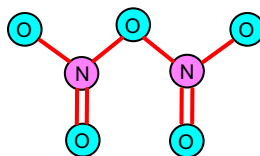
It can also be prepared by the oxidation of  $\text{NO}_2$  with ozone.



As will be discussed later,  $\text{N}_2\text{O}_5$  resembles  $\text{NO}_2^+\text{NO}_3^-$ , and as a result it is suggested that a molecule formally containing  $\text{NO}_2^+$  and another containing nitrate could produce  $\text{N}_2\text{O}_5$ . Thus we observe that a reaction such as the following does indeed take place.



Dinitrogen pentoxide is a white solid that sublimates at  $32^\circ\text{C}$ . The solid is ionic,  $\text{NO}_2^+\text{NO}_3^-$ , with  $\text{NO}_2^+$  being linear ( $D_{\infty h}$ ) as a result of it being a 16-electron triatomic species. The N–O bond length in  $\text{NO}_2^+$  is 115 pm, and at low temperatures, the structure of  $\text{N}_2\text{O}_5$  is



This oxide reacts with water to produce nitric acid.



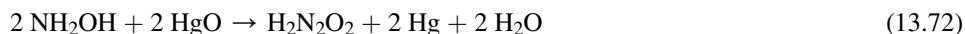
In keeping with the +5 oxidation state of nitrogen in this oxide, it is also a good oxidizing agent.

It appears that the unstable compound  $\text{NO}_3$  exists, particularly in mixtures of  $\text{N}_2\text{O}_5$  and ozone. However, it is not of sufficient importance to discuss here.

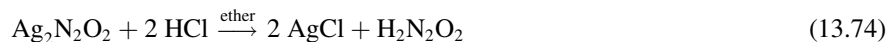
## 13.7 OXYACIDS

### 13.7.1 Hyponitrous Acid, $\text{H}_2\text{N}_2\text{O}_2$

Hyponitrous acid is produced by the reactions



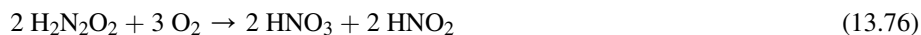
If a solution of the silver salt is treated with HCl dissolved in ether, a solution containing the acid  $\text{H}_2\text{N}_2\text{O}_2$  is obtained.



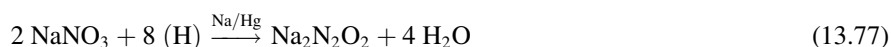
Evaporation of the ether produces solid  $\text{H}_2\text{N}_2\text{O}_2$ , but the acid is so unstable that it explodes readily. In water, the acid decomposes as shown in the following equation:



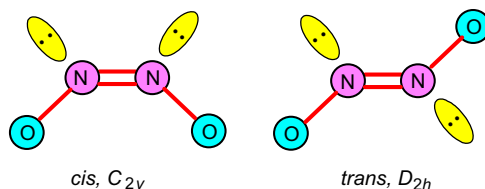
Although  $\text{N}_2\text{O}$  is formally the anhydride of  $\text{H}_2\text{N}_2\text{O}_2$ , the acid cannot be prepared from  $\text{N}_2\text{O}$  and water. In this regard, it is similar to  $\text{CO}$ , which is the anhydride of formic acid,  $\text{HCOOH}$ . Hyponitrous acid is oxidized in air to produce nitric and nitrous acids.



Reduction of a nitrate or nitrite by sodium amalgam in the presence of water has been used to prepare hyponitrite salts.



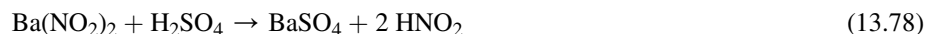
The  $\text{N}_2\text{O}_2^{2-}$  ion exists in two forms,



The *trans* form is more stable, and it is produced by means of the reactions described above.

### 13.7.2 Nitrous Acid, $\text{HNO}_2$

This acid has as its anhydride  $\text{N}_2\text{O}_3$ . The pure acid is unstable so salts are generally employed or a dilute solution of  $\text{HNO}_2$  is generated as needed. A convenient preparation of an aqueous solution containing  $\text{HNO}_2$  is



because the  $\text{BaSO}_4$  produced is a solid that can be easily separated. Heating alkali metal nitrates produces the nitrites.



Aqueous solutions of the acid decompose at room temperature in a disproportionation reaction that can be shown as follows:

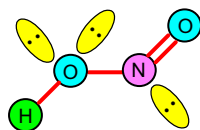


In hot solutions the reaction can be represented by the equation

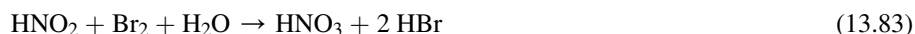




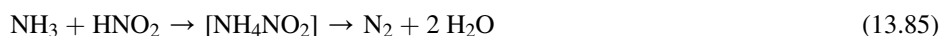
There are several possible structures for  $\text{HNO}_2$ , but the most stable is



Nitrous acid is a weak acid with  $K_a = 4.5 \times 10^{-4}$  that can act as either an oxidizing or reducing agent as illustrated in the following equations.

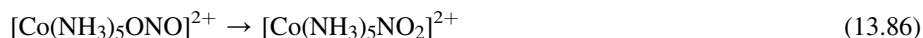


The acid reacts with ammonia to produce  $\text{N}_2$ .



Nitrites are very toxic and in one instance in Chicago, some people died after eating food that had inadvertently been seasoned with sodium nitrite. In that case,  $\text{NaNO}_2$  had accidentally been used to fill a salt shaker. Some nitrites are also explosive.

The  $\text{NO}_2^-$  ion is a good coordinating group and many nitrite complexes are known. It can bind to metals to give  $\text{M}-\text{ONO}$  or  $\text{M}-\text{NO}_2$  linkages, and it can also bridge between two metal centers ( $\text{M}-\text{ONO}-\text{M}$ ). The first known case of linkage isomerization involved the ions  $[\text{Co}(\text{NH}_3)_5\text{NO}_2]^{2+}$  and  $[\text{Co}(\text{NH}_3)_5\text{ONO}]^{2+}$ , which were studied by S. M. Jørgensen in the 1890s. Of these two isomers, the one having the  $\text{Co}-\text{NO}_2$  linkage is more stable, and the reaction

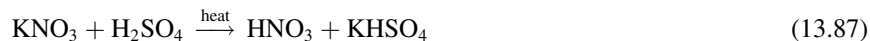


takes place both in solution and in the solid state.

### 13.7.3 Nitric Acid, $\text{HNO}_3$

This is the most important of the acids containing nitrogen and it is used in enormous quantities ( $\sim 19$  billion lbs/year). It is used in the manufacture of explosives, propellants, fertilizers, organic nitro compounds, dyes, plastics, etc. It is a strong acid, because  $b = 2$  in the formula  $(\text{HO})_a\text{XO}_b$ , and is a strong oxidizing agent.

Nitric acid has been known for hundreds of years. In 1650, J. R. Glauber prepared it by the reaction

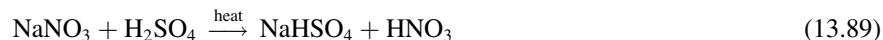


and this method is still used. However, it appears that the acid had been known earlier than this. The pure acid has a density of  $1.55 \text{ g cm}^{-3}$ , a m.p. of  $-41.6^\circ\text{C}$ , and a b.p. of  $82.6^\circ\text{C}$ . It forms a constant boiling mixture (b.p.  $120.5^\circ\text{C}$ ) with water that contains 68%  $\text{HNO}_3$  and has a density of  $1.41 \text{ g cm}^{-3}$ . It forms several well-defined hydrates such as  $\text{HNO}_3 \cdot \text{H}_2\text{O}$ ,  $\text{HNO}_3 \cdot 2\text{H}_2\text{O}$ , etc. The concentrated acid usually has a yellow-brown color because of the presence of  $\text{NO}_2$  that results from slight decomposition as shown by the equation



Extensive hydrogen bonding occurs in the pure acid and in concentrated aqueous solutions.

Nitrates are found in Chile and nitric acid was formerly prepared by the reaction

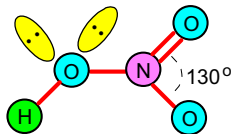


However, the oxidation of ammonia by the *Ostwald process*



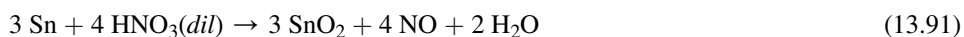
is the basis for the modern production method. Other reactions in the process of converting NO to HNO<sub>3</sub> have been described earlier (see Section 13.4.1).

The HNO<sub>3</sub> molecule has the structure shown below.

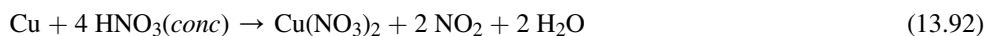


The nitrate ion is planar and has  $D_{3h}$  symmetry (see Chapter 3).

Nitric acid is a strong acid and a strong oxidizing agent so it attacks most metals. In dilute solutions of HNO<sub>3</sub>, NO is produced as the reduction product.



whereas in concentrated HNO<sub>3</sub>, NO<sub>2</sub> is the reduction product.

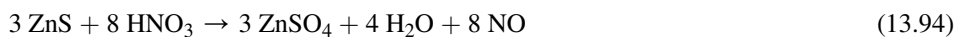


Some of these reactions in which nitric acid reacts as an oxidizing agent have been shown earlier in this chapter. However, some metals such as aluminum form an oxide layer on the surface so that they become passive to further action.

The acid also oxidizes sulfur,



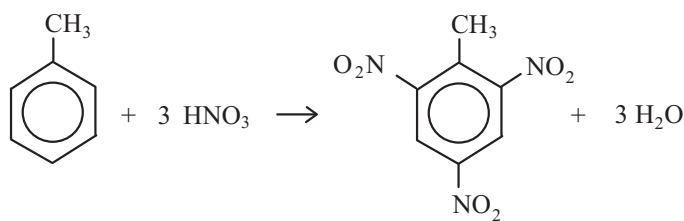
and it oxidizes sulfides to sulfates.



A mixture of one volume of concentrated HNO<sub>3</sub> with three volumes of concentrated HCl is known as *aqua regia*, a solution that will dissolve even gold and platinum.

Many nitrates are important chemicals. For example, black powder (also known as “gunpowder”) has been used for centuries, and it is a mixture containing approximately 75% KNO<sub>3</sub>, 15% C, and 10% S. The mixture is processed while wet, made into flakes, and then dried. Except for use in muzzle loading firearms, it has largely been replaced by smokeless powder that is based on nitrocellulose containing small amounts of certain additives.

The nitration of toluene produces the explosive known as TNT (trinitrotoluene).



2,4,6-trinitrotoluene

(13.95)

This explosive is remarkably stable to shock and it requires a powerful detonator to initiate the explosion. In the explosives industry, precise control of temperature, mixing time, concentrations, heating and cooling rates, etc., is maintained. Making these materials safely requires sophisticated equipment and technology to carry out the process, even though the chemistry may appear simple. Under other conditions, some 2,3,5-trinitrotoluene, 3,5,6-trinitrotoluene, and 2,4,5-trinitrotoluene are produced and they are decidedly less stable than 2,4,6-TNT. A mixture of explosives is only as stable as its least stable component! Without the sophisticated equipment and the knowledge that comes from specialized experience, no one should work with these materials.

## 13.8 NITROGEN IN THE ENVIRONMENT

The atmosphere of the earth is composed of approximately 78% N<sub>2</sub>, making the atmosphere a large reservoir of nitrogen. The element nitrogen is incorporated in living systems in amino acids, proteins, and DNA and is necessary in a large number of biological processes. Because the bond in the N<sub>2</sub> molecule is very strong (946 kJ mol<sup>-1</sup>), the N<sub>2</sub> molecule is very stable and cannot be used directly in these processes. It must be converted into other forms in order for plants and animals to be able to incorporate nitrogen. This constant process of conversion, incorporation, and release of nitrogen-containing compounds into the air and soil is referred to as the *nitrogen cycle*.

### 13.8.1 Nitrogen Fixation

In the soil, certain bacteria contain the *nitrogenase* enzyme that allows atmospheric nitrogen to be converted to ammonia in a process known as *nitrogen fixation*.



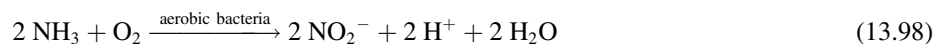
This process occurs best at a temperature of approximately 20–30 °C. The Haber process for the industrial production of ammonia described earlier in this chapter uses the same basic chemical reaction, but high temperatures (approximately 450 °C), high pressures (300 atm), and a catalyst are required in that process.

### 13.8.2 Nitrification

The ammonia produced by nitrogen fixation can be converted to ammonium ions by reaction with water



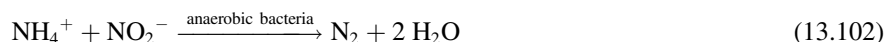
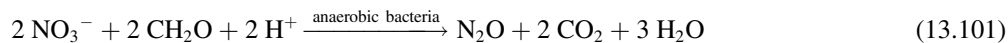
and both ammonia and ammonium ions can then be converted by bacteria to nitrate and nitrate ions in a process known as *nitrification*.



The nitrates thus formed can be incorporated into amino acids and proteins. Rates of nitrification reactions are controlled by the concentrations of ammonium ions, ammonia, and O<sub>2</sub> molecules. Moisture content, pH, and temperature also affect the rates of nitrification.

### 13.8.3 Denitrification

The nitrogen cycle is completed when nitrates in the soil are converted back to nitrous oxide and then nitrogen. This process is known as *denitrification* and representative reactions are shown in [Eqs \(13.101\) and \(13.102\)](#).



Denitrification processes can be used in the treatment of wastewater that has high nitrate levels. These reactions are inhibited by the presence of oxygen and thus reactions such as these occur more quickly when oxygen levels in the soil are low. Rates of denitrification are also controlled by availability of carbon-containing molecules, nitrate levels, temperature, pH, and moisture content.

### 13.8.4 Human Influences in the Nitrogen Cycle

Human activities have disrupted the nitrogen cycle and led to an increase in the amount of reactive nitrogen in the environment. This increase has primarily resulted from reactions that occur in internal combustion engines and from the use of fertilizers.

Because of the high temperatures inside an internal combustion engine,  $N_2$  in the atmosphere can be converted to more reactive forms of nitrogen through the reactions



These processes can also occur naturally during lightning storms. Reactions of  $NO_2$  with water can lead to acid rain.



In addition, the overuse of fertilizers in the form of ammonia or ammonium nitrate to increase the amount of nitrogen in the soil can lead to leaching of nitrogen-containing compounds into groundwater or runoff into bodies of water. The availability of nitrogen in water supplies can cause eutrophication to occur in which phytoplankton grow and produce organic matter. Bacteria then break down the organic matter and in the process consume oxygen. This produces “dead zones” in some waterways in which plants and animals cannot survive because of the lack of  $O_2$ .

Also, increased levels of nitrates in water supplies can be toxic to infants, through a process known as *anoxia* or blue-baby syndrome. In babies, nitrates are converted to nitrites that can bind to hemoglobin to form methemoglobin, which cannot transport  $O_2$ .

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## PROBLEMS

- Starting with  $N_2$ , describe completely the commercial synthesis of nitric acid.
- Explain why numerous nitrogen compounds are explosive.
- Explain why the  $NF_5$  molecule does not exist (although  $NF_4^+F^-$  is known) whereas  $PF_5$  exists as a molecule.
- Draw structures for the following species:
  - Diimine
  - Hydrazine
  - Hyponitrous acid
  - The cyanamide ion
  - Hydroxylamine
- Write the series of balanced equation for the preparation of hydrazine.
- Explain why  $NH_3$  is more polar than  $NCl_3$ .
- Complete and balance the following.
  - $NH_4NO_2 \xrightarrow{\text{heat}}$
  - $Mg + N_2O \rightarrow$
  - $Zn + HNO_3(\text{dil}) \rightarrow$
  - $N_2O_3 + NaOH \rightarrow$
  - $NO_2 + H_2O \rightarrow$
- When a burning strip of magnesium is thrust into a bottle containing  $N_2O$ , it continues to burn. Write the equation for the reaction and explain how it lends support for the structure of  $N_2O$  being what it is known to be.
- Both hyponitrous acid and nitroamide (also known as nitramide) have the formula  $H_2N_2O_2$ . Draw the structures for these molecules and explain any difference in acid–base properties.

10. Draw the structure for nitrosyl chloride. After you have drawn the correct structure, draw structures showing how the atoms would not be arranged and explain why they would not be arranged in that order. What symmetry elements does the nitrosyl chloride possess?
11. Some metals do not react extensively with nitric acid. What is the reason for this behavior?
12. There is only a slight difference in the N–O bond lengths in the FNO, ClNO, and BrNO molecules. Predict which would have the shortest N–O bond length. Which would have the longest N–O bond length? Explain your predictions.
13. Complete the following equations for reactions that take place in liquid ammonia.
- (a)  $\text{Li}_3\text{N} + \text{NH}_4\text{Cl} \rightarrow$
  - (b)  $\text{CaNH} + \text{NH}_3 \rightarrow$
  - (c)  $\text{NaH} + \text{NH}_3 \rightarrow$
  - (d)  $\text{KNH}_2 + \text{NH}_4\text{F} \rightarrow$
  - (e)  $\text{AgCl} + \text{NaNO}_3 \rightarrow$
14. (a) Write the equation for the autoionization of liquid  $\text{N}_2\text{O}_4$ .  
(b) Write the formula for a substance that would be an acid in liquid  $\text{N}_2\text{O}_4$ .  
(c) Write the formula for a substance what would be a base in liquid  $\text{N}_2\text{O}_4$ .  
(d) Write the neutralization reaction for the acid in (b) with the base in (c).
15. The base constants for hydrazine are  $K_{b1} = 8.5 \times 10^{-7}$  and  $K_{b2} = 8.9 \times 10^{-16}$ . Explain this large difference between the  $K_b$  values.
16. The bond angles in FNO, ClNO, and BrNO are  $110^\circ$ ,  $113^\circ$ , and  $117^\circ$ , respectively. Explain this trend.
17. The barrier for rotation around the N–N bond in hydrazine is much larger than for rotation around the C–C bond in ethane. Explain this difference.
18. A molecule having the empirical formula  $\text{N}_4\text{O}$  is nitrosyl azide. Draw the structure of this molecule and describe the bonding.
19. In most compounds that contain N=O bonds, the length of the bond is approximately 120 pm. Explain why the bond length in  $\text{NO}^+$  and gaseous NO differ from that value.
20. Complete and balance the following.
- (a)  $\text{CaCN}_2 + \text{H}_2\text{O} \rightarrow$
  - (b)  $\text{NH}_4\text{NO}_3 \xrightarrow{\text{heat}}$
  - (c)  $\text{NBr}_3 + \text{H}_2\text{O} \rightarrow$
  - (d)  $\text{SeF}_4 + \text{N}_2\text{F}_4 \rightarrow$
  - (e)  $\text{PF}_5 + \text{N}_2\text{F}_2 \rightarrow$
21. Account for the fact that the N–N bond length in  $\text{N}_2\text{O}_3$  is 186 pm although that in  $\text{N}_2\text{H}_4$  is 145 pm.
22. Explain why the bonds between the nitrogen and two of the oxygen atoms in the nitric acid molecule are slightly shorter than to the third.
23. How would you expect the N–N bond length in  $\text{N}_2\text{H}_5^+$  to compare to that in  $\text{N}_2\text{H}_6^{2+}$ ? Explain your answer.

# Phosphorus, Arsenic, Antimony, and Bismuth

The elements constituting Group VA in the periodic table have an enormous range of chemical properties. As a general trend, there is an increase in metallic character progressing downward in the group. Each of the elements has possible oxidation states that range from  $-3$  to  $+5$ , although part of the range is not important for some of the elements. These elements also have extensive organic or organometallic chemistry depending on the electronegativity of the element.

Although all of the elements are important and are found in many common compounds, phosphorus and its compounds are among the most useful and essential of any element. Consequently, the chemistry of phosphorus is considerably more extensive, and it is covered in greater detail. Much of the chemistry of the other elements can be inferred from their greater metallic character and by comparisons to the analogous phosphorus compounds.

## 14.1 OCCURRENCE

Phosphorus occurs extensively in nature, the most common materials being phosphate rocks and minerals, bones, and teeth. Most important of the phosphate rocks are calcium phosphate,  $\text{Ca}_3(\text{PO}_4)_2$ , *apatite*,  $\text{Ca}_5(\text{PO}_4)_3\text{OH}$ , *fluoroapatite*,  $\text{Ca}_5(\text{PO}_4)_3\text{F}$ , and *chloroapatite*,  $\text{Ca}_5(\text{PO}_4)_3\text{Cl}$ . The element was first obtained by Brandt, and its name is derived from two Greek words meaning “light” and “I bear” because white phosphorus glows in the dark as a result of slow oxidation (phosphorescence).

Arsenic is found primarily as sulfides in *orpiment*,  $\text{As}_2\text{S}_3$ , *realgar*,  $\text{As}_4\text{S}_4$ , and *arsenopyrite*,  $\text{FeAsS}$ , but it also occurs as *arsenolite*,  $\text{As}_4\text{O}_6$ . Arsenic compounds have been known since antiquity, and *orpiment*, which is yellow, was probably used as a yellow pigment.

The most important mineral containing antimony is *stibnite*,  $\text{Sb}_2\text{S}_3$ . Antimony sulfide was used as a dark material for painting around the eyes by women in ancient Egypt and Persia. Other minerals containing arsenic include *ullmanite*,  $\text{NiSbS}$ , *tetrahedrite*,  $\text{Cu}_3\text{SbS}_3$ , and a number of other complex sulfides.

Bismuth is found in some locations in the free state. Generally, it is found as *bismite*,  $\text{Bi}_2\text{O}_3$ , *bismuth glance*,  $\text{Bi}_2\text{S}_3$ , or in a few other minerals. The name of the element is thought to be derived from a German word meaning “white matter.”

## 14.2 PREPARATION AND PROPERTIES OF THE ELEMENTS

Phosphorus is obtained commercially from rocks bearing the phosphate minerals described above. In the process, crushed phosphate rock is treated with carbon and silica in an electric furnace at  $1200\text{--}1400\text{ }^\circ\text{C}$ . Under these conditions, the phosphorus is distilled out.



At temperatures below  $800\text{ }^\circ\text{C}$ , elemental phosphorus exists as  $\text{P}_4$  molecules having a tetrahedral structure. At the temperature used in the preparation of phosphorus, some of the molecules dissociate to  $\text{P}_2$ .

There are several allotropic forms of elemental phosphorus, the most common being the white, red, and black forms. Red phosphorus, which itself includes several forms, is obtained by heating the white form at  $400\text{ }^\circ\text{C}$  for several hours. An amorphous red form may also be prepared by subjecting white phosphorus to ultraviolet radiation. In the thermal transformation, several substances function as catalysts, e.g., iodine, sodium, and sulfur. Black phosphorus appears to consist of four different forms. These are obtained by the application of heat and pressure to the white form. The major uses of elemental phosphorus involve the production of phosphoric acid and other chemicals. Red phosphorus is used in making matches and white phosphorus has had extensive use in making incendiary devices. Several of the important classes of phosphorus compounds will be discussed in later sections.

Arsenic is usually obtained by the reduction of its oxide with carbon.



The oxide is obtained from the sulfide by roasting it in air. The stable form of arsenic is the gray or metallic form although other forms are known. Yellow arsenic is obtained by cooling the vapor rapidly and an orthorhombic form is obtained when the vapor is condensed in the presence of mercury. Arsenic is used in the production of a variety of insecticides and herbicides, and in alloys with copper and lead. Some arsenic compounds are important medicinal compounds and a number of pigments contain arsenic compounds. The surface tension of lead is increased by dissolving a small amount of arsenic in it. This allows droplets of molten lead to assume a spherical shape, and this fact is utilized in the production of lead shot.

Antimony is obtained by reduction of the sulfide ore with iron.



Reduction of antimony oxide (obtained by roasting the sulfide ores) with carbon is also employed. The most stable form of the element has a rhombohedral structure, although at high pressure this form converts to others having cubic and hexagonal close packing structures. Several amorphous forms are also known. Antimony has the property of hardening lead when alloyed with it. Thus, many uses of lead, e.g., in automobile batteries, are based on the fact that antimony hardens and strengthens the lead. Such alloys expand on cooling and give a sharp casting without shrinking away from the mold. For this reason, antimony has historically been used in type metal.

The elements in Group VA are frequently found in structures in which there are clusters of atoms (e.g.,  $\text{P}_4$ ). However, many polyanions are also known with the most common being those such as  $\text{P}_7^{3-}$ ,  $\text{As}_7^{3-}$ , etc. Some of the chemistry of such species was discussed in Section 8.9 in keeping with the most common cations in the compounds being those of Group IA or IIA metals.

### 14.3 HYDRIDES

The elements in Group VA of the periodic table form several binary compounds with hydrogen. Some of these are analogous to the hydrogen compounds formed by nitrogen ( $\text{NH}_3$ ,  $\text{N}_2\text{H}_4$ , and  $\text{HN}=\text{NH}$ ). However, one considerable difference is that the hydrides of the heavier elements are much less basic. In nitrogen compounds, the unshared pair constitutes a basic site that is hard (see Chapter 6). Therefore, toward proton donors the nitrogen compounds are distinctly basic. In the case of  $\text{PH}_3$ ,  $\text{PR}_3$ ,  $\text{AsH}_3$ ,  $\text{AsR}_3$ , etc., the unshared pair of electrons resides in a rather large orbital. Consequently, these molecules behave as soft bases, and they do not interact well with protons. Accordingly,  $\text{PH}_3$  is a much weaker base toward  $\text{H}^+$  than is  $\text{NH}_3$ , and the reaction of phosphine with water is so slight that the solutions are essentially neutral. In liquid ammonia, phosphine reacts as an acid to produce  $\text{NH}_4^+\text{PH}_2^-$ . Phosphonium salts are generally stable only when the proton donor is a strong acid and the anion is large so that there is a rather close match in the size of anion and cation. A typical case is illustrated by the reaction with HI.



Arsenic, antimony, and bismuth do not form stable compounds containing  $\text{AsH}_4^+$ ,  $\text{SbH}_4^+$ , and  $\text{BiH}_4^+$  ions.

On the other hand, toward soft electron pair acceptors such as  $\text{Pt}^{2+}$ ,  $\text{Ag}^+$ ,  $\text{Ir}^+$ , etc., the phosphines are stronger Lewis bases than are  $\text{NH}_3$  and amines. In other words, phosphines and arsines are better ligands toward Class B metals than are amines (see Chapter 6). The stable complexes are those with second and third row transition metals in low oxidation states.

Due to the difference in size between the atomic orbitals that form the bonds,  $\text{PH}_3$  is less stable than  $\text{NH}_3$ . In fact, the stability of the  $\text{EH}_3$  compounds (named as phosphine, arsine, stibine, and bismuthine) decreases as one goes down the group. Similar trends are seen for hydrogen compounds of the Group IVA, VIA, and VIIA elements (see Chapters 12, 16, and 17). The physical properties of the hydrogen compounds of the Group VA elements are shown in Table 14.1 with the properties of  $\text{NH}_3$  included for comparison.

The bond angles in the  $\text{EH}_3$  molecules reveal some insight into the nature of the bonding. For example in  $\text{NH}_3$ , the bond angle is about that expected for  $sp^3$  hybridization of the nitrogen orbitals with some decrease caused by the unshared pair of electrons on the nitrogen atom. The bond angles in the hydrides of the remainder of the Group VA elements indicate almost pure  $p$  orbital bonding by the central atom. An explanation for this is provided by considering the fact that orbitals of similar size overlap best. In the case of the nitrogen atom, the  $sp^3$  orbitals are not too much larger than the hydrogen  $1s$  orbital for the overlap to be effective in producing stronger bonds. For the larger atoms, the  $p$  orbitals are already larger

**TABLE 14.1** Properties of the Hydrogen Compounds of the Group VA Elements

	NH <sub>3</sub>	PH <sub>3</sub>	AsH <sub>3</sub>	SbH <sub>3</sub>	BiH <sub>3</sub>
m.p., °C	-77.7	-133.8	-117	-88	—
b.p., °C	-33.4	-87.8	-62.5	-18.4	17
$\Delta H_f^\circ$ , kJ mol <sup>-1</sup>	-46.11	9.58	66.44	145.10	277.8
H—E—H angle, deg	107.1	93.7	91.8	91.3	—
$\mu$ , D	1.46	0.55	0.22	0.12	—

than the hydrogen 1s orbital and increasing their size by making  $sp^3$  hybrids decreases the effectiveness of overlap. Therefore, stronger bonds form with the essentially unhybridized  $p$  orbital on the central atom and the bond angles are close to the 90° expected if the central atom were using pure  $p$  orbitals. This trend is similar to that observed for the hydrogen compounds of the Group VIA elements (see Chapter 16). Also, when the central atom is phosphorus rather than nitrogen, the bonding pairs of electrons are farther apart than when the central atom is nitrogen. When nitrogen is the central atom, the bonding pairs are drawn inward so a greater amount of  $s$  character in the hybrid orbital ( $sp^3$  rather than  $p^3$ ) increases the bond angle which leads to less repulsion.

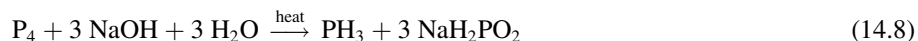
In addition to the  $EH_3$  compounds, a few others such as  $P_2H_4$ ,  $As_2H_4$ ,  $Sb_2H_4$ , are also known. All of the hydrogen compounds are extremely toxic. These compounds are comparatively unstable, and in fact,  $P_2H_4$  is spontaneously flammable. Phosphine,  $PH_3$ , is a reducing agent, and it is easily oxidized by burning in air.



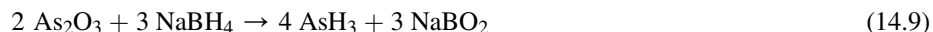
A general preparation of the trihydride compounds involves the formation of a metal compound with the Group VA element and the reaction of that compound with water or an acid. The process can be illustrated by the following equations.



Phosphine can also be prepared by the reaction of elemental phosphorus with a hot solution of a strong base.



Arsine can be prepared by the reaction of  $As_2O_3$  with  $NaBH_4$  with the overall reaction being represented as follows.



It can also be prepared by the reaction of arsenic with an active metal to give an arsenide that will react with water or an acid to give arsine.



The hydrides of the heavier members of Group VA are easily decomposed. For example, the decomposition of arsine serves as a basis for the Marsh test for arsenic in which an arsenic mirror forms when arsine is passed through a heated tube.

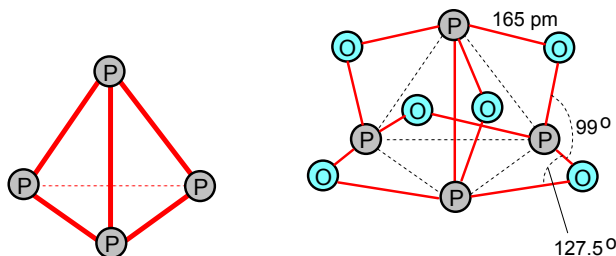


The trihydrides of the Group VA elements are all extremely toxic gases.

## 14.4 OXIDES

Although other oxides are known, only two series of oxides are of any great importance for the elements in Group VA. Consequently, the discussion will be limited to these two series, those in which the oxidation states are +3 and +5.





**FIGURE 14.1** The structure of the  $P_4O_6$  molecule is based on the  $P_4$  tetrahedron with oxygen bridges between the phosphorus atoms.

### 14.4.1 The +3 Oxides

Phosphorus can be oxidized to yield phosphorus(III) oxide when the amount of oxygen is controlled.



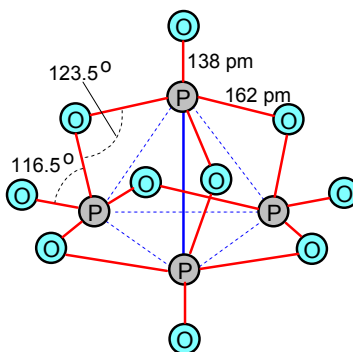
In the  $P_4O_6$  molecule, the tetrahedral arrangement of phosphorus atoms is maintained so that the structure of the molecule is as shown in Figure 14.1. This compound melts at  $23.9^\circ\text{C}$  and boils at  $175.4^\circ\text{C}$ . Chemically, this oxide is the anhydride of phosphorous acid,  $H_3PO_3$ . In the reaction with cold water,  $P_4O_6$  does produce that acid. In hot water, disproportionation occurs and phosphine, phosphorus, and phosphoric acid are produced. At temperatures much above its boiling point,  $P_4O_6$  decomposes into phosphorus and an oxide that can be described by the formula  $P_nO_{2n}$ .

The oxides  $As_4O_6$  and  $Sb_4O_6$  are formed by burning the metals in air. The structures of these oxides are similar to that of  $P_4O_6$ . Although As and Sb form +5 oxides,  $Bi_2O_3$  is the only common oxide of bismuth. Several forms of the solid oxides are known that involve oxygen atoms in bridging positions. Particularly in the cases of  $As_2O_3$ ,  $Sb_2O_3$ , and  $Bi_2O_3$ , the oxides are the starting point for synthesis of many other compounds of these elements.

### 14.4.2 The +5 Oxides

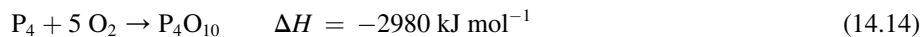
The +5 oxides of arsenic, antimony, and bismuth have not been well characterized and comparatively little is known about such materials. On the other hand,  $P_4O_{10}$  is an important compound, and it deserves to be discussed in some detail. The actual molecular formula is  $P_4O_{10}$ , not  $P_2O_5$  although the compound was considered to be  $P_2O_5$  for many years. The structure of the  $P_4O_{10}$  molecule shown in Figure 14.2 is derived from the tetrahedral structure of  $P_4$  with bridging oxygen atoms along each edge of the tetrahedron and an oxygen atom in a terminal position on each phosphorus atom.

Three crystalline forms of  $P_4O_{10}$  are known. The most common of these is the hexagonal or *H*-form, and the others are orthorhombic forms known as the *O*- and *O'*-forms. Heating the *H*-form at  $400^\circ\text{C}$  for 2 h produces the *O*-form and heating the *H*-form at  $450^\circ\text{C}$  for 24 h produces the *O'*-form. Both of the orthorhombic forms are less reactive than the *H*-form.



**FIGURE 14.2** The structure of the  $P_4O_{10}$  molecule.

Because the actual formula for phosphorus(V) oxide is  $P_4O_{10}$ , the compound can also be named as tetraphosphorus decoxide. It is the anhydride of the phosphoric acids and as such it is produced as the first step in the manufacture of  $H_3PO_4$ . It is readily obtained by burning elemental phosphorus.



The compound is a powerful dehydrating agent that is sometimes used as a desiccant, and the reaction with water can be shown as

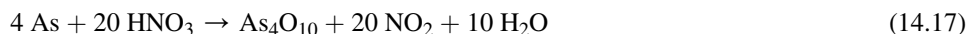


A large number of organic phosphates are prepared by the reactions of  $P_4O_{10}$  with alcohols.



Monoalkyl and dialkyl derivatives are also obtained in these reactions. Alkyl phosphates are used as acid catalysts, lubricants, and intermediates in preparing flame-proofing compounds.

Both arsenic(V) oxide and antimony(V) oxide are prepared by the reaction of the elements with concentrated nitric acid.



However, neither the arsenic oxide nor antimony oxide is very important. An oxide of antimony having the formula  $Sb_2O_4$  is also known. Although this compound formally contains antimony(IV), it is known that it actually contains equal numbers of Sb(III) and Sb(V) atoms. Antimony also shows this type of behavior in forming bridged species such as  $Sb_2Cl_{10}^{2-}$  that contain equal numbers of Sb(III) and Sb(V) rather than Sb(IV).

## 14.5 SULFIDES

At temperatures above 100 °C, phosphorus and sulfur react to produce several binary compounds. Of these,  $P_4S_{10}$ ,  $P_4S_7$ ,  $P_4S_5$ , and  $P_4S_3$  have been the most thoroughly studied. The structure of  $P_4S_{10}$  is similar to that of  $P_4O_{10}$  except for sulfur atoms replacing the oxygen atoms. The structures of these compounds contain some P–P bonds and some P–S–P bridges. These structures can be considered as being derived from the  $P_4$  tetrahedron by insertion of bridging sulfur atoms.

Tetraphosphorus trisulfide, also known as phosphorus sesquisulfide,  $P_4S_3$ , can be obtained by heating a stoichiometric mixture of  $P_4$  and sulfur at 180 °C under an inert atmosphere. The compound melts at 174 °C and it is soluble in toluene, carbon disulfide, and benzene. It is  $P_4S_3$  that is used with potassium chlorate, sulfur, and lead dioxide in matches.

Tetraphosphorus pentasulfide,  $P_4S_5$ , is obtained by the reaction of a solution of sulfur in  $CS_2$  with  $P_4S_3$  in the presence of  $I_2$  as a catalyst. Tetraphosphorus heptasulfide,  $P_4S_7$ , is prepared by heating phosphorus and sulfur in a sealed tube. Neither  $P_4S_5$  nor  $P_4S_7$  have important uses.  $P_4S_{10}$  is prepared by reaction of the elements in stoichiometric amounts.



The reaction of  $P_4S_{10}$  with water can be represented as follows.



Heating  $P_4O_6$  with sulfur produces an oxysulfide having the formula  $P_4O_6S_4$  that has a structure similar to that of  $P_4O_{10}$  except the terminal (nonbridging) oxygen atoms are replaced by sulfur atoms. A few other mixed sulfur–oxygen compounds of phosphorus are known.

Several sulfides of arsenic, antimony, and bismuth are known. Some of these, e.g.,  $As_4S_4$ ,  $As_2S_3$ ,  $Sb_2S_3$ , and  $Bi_2S_3$ , are the important minerals containing these elements. Some of the sulfides can be precipitated from aqueous solutions owing to the insolubility of sulfides of As(III), As(V), Sb(III), and Bi(III). As in the case of phosphorus, arsenic also forms a sesquisulfide,  $As_4S_3$ . Several of these brightly colored sulfide compounds are used as pigments, and some behave as semiconductors. This is also true of some of the selenides and tellurides of arsenic, antimony, and bismuth.

## 14.6 HALIDES

The halogen compounds of the Group VA elements are reactive compounds that generally react by breaking the bonds to the halogen. However, as a result of having an unshared pair of electrons, they can also function as Lewis bases, and they are useful as starting materials for preparing a wide variety of compounds, both inorganic and organic.

### 14.6.1 Halides of the Type E<sub>2</sub>X<sub>4</sub>

A number of halogen compounds of the Group VA elements are known that have the formula E<sub>2</sub>X<sub>4</sub>. None of the compounds is of any great importance, and we will describe only the phosphorus compounds here. The compounds having this formula retain one E–E bond. When E is phosphorus, the fluoride, chloride, and iodide compounds are known. The preparation of P<sub>2</sub>F<sub>4</sub> is carried out by the reaction of PF<sub>2</sub>I with mercury, which removes iodine leading to a coupling reaction.



The chloride compound is obtained by electric discharge in a mixture of PCl<sub>3</sub> and H<sub>2</sub>. Preparation of P<sub>2</sub>I<sub>4</sub> can be carried out by the reaction of I<sub>2</sub> with a carbon disulfide solution of white phosphorus.

### 14.6.2 Trihalides

All of the trihalides of the Group VA elements are known. In principle, the direct action of the appropriate halogen on the elements leads to the formation of the trihalides. However, such reactions may not always be the best preparative methods. The fluorides are prepared by the following reactions.



Bismuth trifluoride is only slightly soluble and it precipitates from solutions containing Bi<sup>3+</sup> when an excess of F<sup>−</sup> is added.

The chlorides of the Group VA elements are prepared in the following ways. Phosphorus trichloride is obtained by the reaction of white phosphorus with chlorine using an excess of phosphorus.



The trichlorides of As, Sb, and Bi are produced when As<sub>4</sub>O<sub>6</sub>, Sb<sub>2</sub>S<sub>3</sub>, and Bi<sub>2</sub>O<sub>3</sub> react with concentrated HCl.

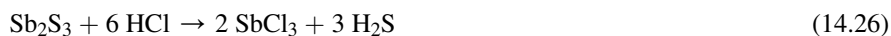


Table 14.2 shows some of the properties of the trihalides of the Group VA elements.

Several trends in the data shown in Table 14.2 are of interest. For example, the trihalides of phosphorus and arsenic can be considered as covalent molecules. As a result, the intermolecular forces are dipole–dipole and London forces that are weak. Therefore, the melting and boiling points increase with molecular weight as expected. The trifluorides of antimony and bismuth are essentially ionic compounds and the melting points are much higher than those of the halogen derivatives that are more covalent.

A second trend that is interesting is shown by the bond angles of the molecules. For example, the bond angles for the PX<sub>3</sub> molecules are all in the 101° to 104° range that is indicative of sp<sup>3</sup> orbitals being used on the phosphorus atom. For the SbX<sub>3</sub> molecules, the bond angle is only about 90° for SbF<sub>3</sub>, but it is somewhat larger for the other halides. These data indicate that the orbitals used by Sb in SbF<sub>3</sub> are essentially pure p orbitals, whereas hybrids approaching sp<sup>3</sup> are utilized in the chloride, bromide, and iodide compounds. One explanation for this is that the smaller size of the p orbitals of F do not overlap as effectively with larger sp<sup>3</sup> orbitals on Sb so the energy required for hybridization is not recovered by forming stronger bonds. This trend was also seen in the bond angles of the EH<sub>3</sub> molecules.

**TABLE 14.2** Physical Properties of Trihalides of Group VA Elements

Compound	m.p., °C	b.p., °C	$\mu$ , D	X–E–X Angle
PF <sub>3</sub>	–151.5	–101.8	–	104
PCl <sub>3</sub>	–93.6	76.1	0.56	101
PBr <sub>3</sub>	–41.5	173.2	–	101
PI <sub>3</sub>	61.2	dec.	–	102
AsF <sub>3</sub>	–6.0	62.8	2.67	96.0
AsCl <sub>3</sub>	–16.2	103.2	1.99	98.4
AsBr <sub>3</sub>	31.2	221	1.67	99.7
AsI <sub>3</sub>	140.4	370 d	0.96	100.2
SbF <sub>3</sub>	292	–	–	88
SbCl <sub>3</sub>	73	223	3.78	99.5
SbBr <sub>3</sub>	97	288	3.30	97
SbI <sub>3</sub>	171	401	1.58	99.1
BiF <sub>3</sub>	725	–	–	–
BiCl <sub>3</sub>	233.5	441	–	100
BiBr <sub>3</sub>	219	462	–	100
BiI <sub>3</sub>	409	–	–	–

Trihalides of the Group VA elements are pyramidal ( $C_{3v}$ ) with an unshared pair of electrons on the central atom. Typically, the molecules are Lewis bases and they form acid–base adducts and metal complexes. In accord with the hard–soft interaction principle, these species are better electron pair donors toward soft electron pair acceptors. Therefore, most of the complexes of these  $EX_3$  molecules contain second and third row transition metals or first row metals in low oxidation states.

In addition to the trihalides containing the same halogen, some mixed trihalides are known. For example, there is some exchange when two different trihalides are mixed.



All of the trihalides of the Group VA elements hydrolyze in water, but the rates vary in the order  $P > As > Sb > Bi$ , an order that parallels the decreased tendency toward covalent bonding in the molecules. This is also a manifestation of the increase in metallic character in going down in the group. However, the trihalides do all not give the same type of products as a result of hydrolysis. For example, the phosphorus trihalides react according to the equation



In the case of  $PI_3$ , this reaction is a convenient synthesis of HI (see Chapter 7). The arsenic trihalides hydrolyze in an analogous way. However, the trihalides of antimony (and the analogous bismuth compounds) react according to the equation



Antimonyl chloride, sometimes referred to as antimony oxychloride, is known as a “basic chloride.” It is insoluble in water, but aqueous solutions containing the trihalides can be made if the solution contains a sufficiently high concentration of  $HX$  to drive the system to the left. When such a solution is diluted by the addition of water, the oxychloride reprecipitates.

Many of the other reactions of the trihalides of the Group VA elements are essentially the same for all of the elements in the group. Of all the trihalide compounds,  $PCl_3$  is the most important, and it will be discussed further here. In addition to

the hydrolysis reaction,  $\text{PCl}_3$  reacts with oxidizing agents to produce  $\text{POCl}_3$  (more accurately written as  $\text{OPCl}_3$  because the oxygen atom is bonded to the phosphorus atom).



Oxidizing agents such as  $\text{H}_2\text{O}_2$  and  $\text{OCl}^-$  can be employed in this reaction. Reactions of  $\text{PCl}_3$  with other halogens yield pentahalides that contain two different halogens. For example, a mixed pentahalide can be prepared by the reaction



Phosphorus trichloride reacts with organic compounds including Grignard reagents and metal alkyls to give numerous alkyl derivatives. Typical reactions can be illustrated by the following equations.



The dialkyl and trialkyl compounds can also be obtained by using different ratios of the alkylating agent to  $\text{PCl}_3$ . At high temperatures (600–700 °C),  $\text{PCl}_3$  reacts with benzene to produce phenyl dichlorophosphine.



Although the organic chemistry of phosphorus is extensive, some of the most important compounds are the various phosphite esters. The most convenient synthesis of these compounds is the reaction of  $\text{PCl}_3$  with alcohols. However, the reactions can be carried out in more than one way. The direct reaction of  $\text{PCl}_3$  with an alcohol can be shown as



As in the case of phosphorous acid, the product in this reaction has one hydrogen atom bonded directly to the phosphorus atom. When a base such as an amine (represented as Am in the following equation) is present, the reaction can be shown as



In this process, the production of the phosphite is assisted by the reaction of the HCl as it is produced to form an amine hydrochloride.

### 14.6.3 Pentahalides and Oxyhalides

Sixteen pentahalides could conceivably exist for combinations of P, As, Sb, and Bi with F, Cl, Br, and I. However, several of these possibilities are unknown and others are of very little importance. None of the elements forms a stable pentafluoride, but all of the pentafluorides are known. Phosphorus forms a pentachloride and a pentabromide, whereas antimony forms a pentachloride. As with the discussion of the trihalides, most of the discussion presented will center on the phosphorus compounds.

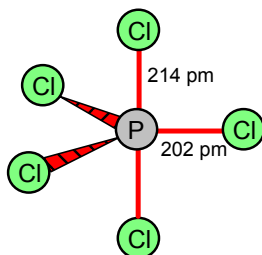
Phosphorus pentachloride can be formed by the reaction of excess chlorine with elemental phosphorus or  $\text{PCl}_3$ .



Phosphorus pentabromide can be prepared by the reaction of  $\text{PBr}_3$  with bromine.



In the gas phase, phosphorus pentachloride and pentabromide have a trigonal bipyramid structure ( $D_{3h}$  symmetry),

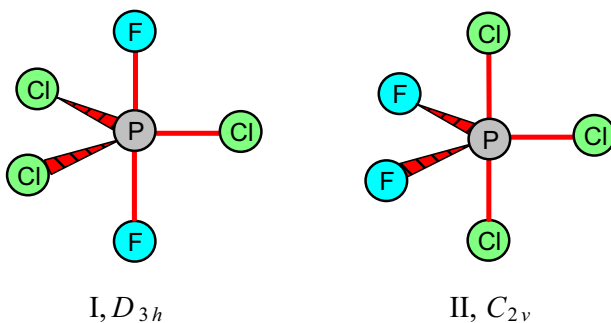


but in the solid phase, both  $\text{PCl}_5$  and  $\text{PBr}_5$  are ionic solids that exist as  $\text{PCl}_4^+\text{PCl}_6^-$  and  $\text{PBr}_4^+\text{Br}_6^-$ , respectively. Solid  $\text{SbCl}_5$  exists as  $\text{SbCl}_4^+\text{Cl}^-$ . In solid  $\text{PCl}_4^+\text{PCl}_6^-$ , the P–Cl bond lengths are 198 pm in the cation and 206 pm in the anion.

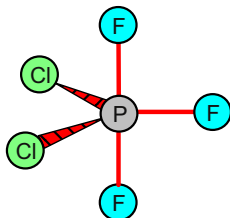
Several mixed halide compounds are known that have formulas, such as  $\text{PCl}_3\text{F}_2$ ,  $\text{PF}_3\text{Cl}_2$ ,  $\text{PF}_3\text{Br}_2$ , etc. These compounds are prepared by adding a halogen to phosphorus trihalides that contain a different halogen, and a typical reaction is



The structures of these compounds are of interest because they are derived from the trigonal bipyramid structure in which there are differences between the axial and equatorial positions. Consider, for example, the molecules  $\text{PCl}_3\text{F}_2$  and  $\text{PF}_3\text{Cl}_2$ . For  $\text{PCl}_3\text{F}_2$  two of the possible structures are



In Chapter 3, the structures of molecules such as  $\text{ClF}_3$  were described as having two unshared pairs of electrons in the equatorial positions. Because this leads to less repulsion, it would be expected that the larger chlorine atoms should occupy the equatorial positions as shown in structure I. This is, in fact, the structure of  $\text{PCl}_3\text{F}_2$  and the molecule has  $D_{3h}$  symmetry. By the same line of reasoning, it would be expected the structure of  $\text{PF}_3\text{Cl}_2$  should be



and the molecule has  $C_{2v}$  symmetry rather than  $D_{3h}$ . However, at temperatures above  $-22^\circ\text{C}$ , the nuclear magnetic resonance spectrum for  $\text{PF}_3\text{Cl}_2$  shows only one peak attributable to fluorine. This indicates that all of the fluorine atoms are in equivalent positions in the structure. At  $-143^\circ\text{C}$ , the spectrum indicates that there are fluorine atoms in two types of environments as is indicated in the structure above. Therefore, at low temperature  $\text{PF}_3\text{Cl}_2$  is “frozen” in the configuration shown above, but at higher temperatures there is sufficient thermal energy available for the molecule to undergo rapid structural changes, which result in the fluorine atoms all appearing to be equivalent. Molecules that have this ability are known as *fluxional* molecules. One mechanism by which this can occur is known as the Berry (after the originator, R. Steven Berry) pseudo rotation in which no bonds are broken.

The pentahalides of Group VA are all strong Lewis acids that interact with electron pair donors to form complexes. This is typified by the formation of halo complexes,



In fact, this tendency is so great that it can result in the generation of unusual cations by fluoride abstraction. For example, the  $\text{ClF}_2^+$  cation can be generated by the reaction

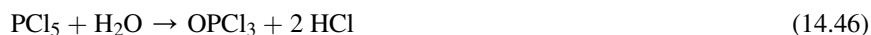


This type of reaction is also discussed in Chapters 6 and 17. Because of their strong Lewis acidity,  $\text{PCl}_5$ ,  $\text{PBr}_5$ ,  $\text{SbCl}_5$ , and  $\text{SbF}_5$  are all effective acid catalysts for reactions such as the Friedel–Crafts reaction.

One of the uses of the pentahalides is in various types of reactions as halogen transfer reagents to both inorganic and organic substrates. For example, the halogenation of  $\text{SO}_2$  by  $\text{PCl}_5$  takes place according to the equation



The oxychlorides can also be obtained by partial hydrolysis, and these compounds are extremely useful intermediates.



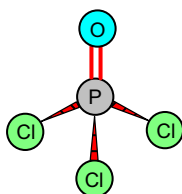
Complete hydrolysis of  $\text{PCl}_5$  results in the formation of phosphoric acid.



Phosphoryl chloride,  $\text{OPCl}_3$ , can also be prepared by the oxidation of  $\text{PCl}_3$  or the reaction of  $\text{P}_4\text{O}_{10}$  with  $\text{PCl}_5$ .



The structure of the  $\text{OPCl}_3$  molecule is



so the molecule possesses  $C_{3v}$  symmetry. This compound still retains three reactive phosphorus–halogen bonds so it undergoes reactions typical of covalent halides. The following equations illustrate these reactions.



Organic phosphates,  $(\text{RO})_3\text{PO}$ , are useful compounds that have a wide range of industrial uses. The chemistry of  $\text{OPCl}_3$  and  $\text{OP}(\text{OR})_3$  has been extensively studied and these solvents usually coordinate to Lewis acids through the terminal oxygen atom.

The sulfur analog,  $\text{SPCl}_3$ , and several mixed oxyhalides such as  $\text{OPCl}_2\text{F}$  and  $\text{OPCl}_2\text{Br}$  are known but they are not widely used compounds. Oxyhalide compounds of arsenic, antimony, and bismuth are of much less importance than are those of phosphorus.

There is a strong tendency of the pentahalides of the heavier members of Group VA to form complexes. The interaction of  $\text{SbCl}_5$  and  $\text{SbCl}_3$  shows that there is a complex formed between the two molecules,



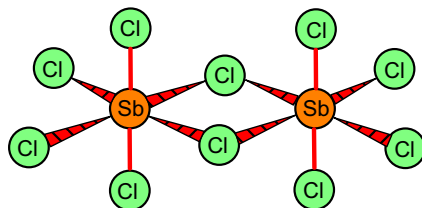
Also, in concentrated  $\text{HCl}$ ,  $\text{SbCl}_5$  is present as a hexachloro complex as a result of the reaction



In aqueous  $\text{HCl}$  solutions, the equilibrium involving the tri- and pentahalides can be shown as



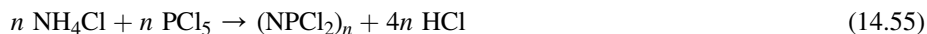
This ion is believed to have a structure in which there are two chloride ion bridges.



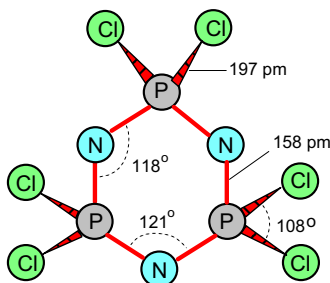
Many other complexes containing species such as  $\text{AsF}_6^-$ ,  $\text{SbF}_6^-$ , and  $\text{PF}_6^-$  are also known.

## 14.7 PHOSPHAZINE (PHOSPHONITRILIC) COMPOUNDS

Although a variety of compounds are known that contain P–N bonds, the *phosphazines* or *phosphonitrilic* compounds are the most interesting. Compounds having both linear and cyclic molecular structures are known in which there are alternating P and N atoms. Historically, the chlorides having the general formula  $(\text{PNCl}_2)_n$  were the first compounds of this type. The most thoroughly studied compounds of this are the triphosphazines,  $(\text{PNX}_2)_3$ . A general preparation involves the reaction of ammonium chloride and phosphorus pentachloride.



This reaction can be carried out in a sealed tube or in a solvent such as  $\text{C}_2\text{H}_2\text{Cl}_4$ ,  $\text{C}_6\text{H}_5\text{Cl}$ , or  $\text{OPCl}_3$ . The cyclic trimer which has the structure

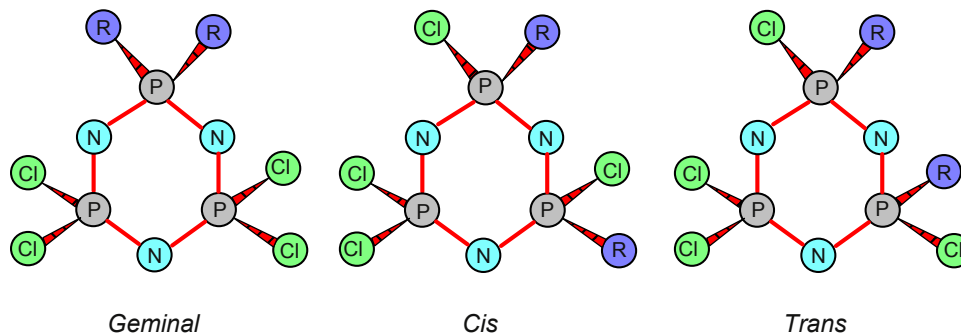


is the most thoroughly studied. In this structure, the P–N bond length is about 158 pm, whereas the usual P–N single bond length is about 175 pm indicating a great deal of multiple bonding. The tetramer,  $(\text{NPCl}_2)_4$ , and the pentamer,  $(\text{NPCl}_2)_5$ , exist as puckered and planar rings, respectively.

When heated at 250 °C,  $(\text{NPCl}_2)_3$  polymerizes to give materials having as many as 15,000 units.



A large number of derivatives of the phosphazine compounds can be prepared by means of the reactions at the P–Cl bonds. Hydrolysis reactions produce P–OH bonds that can then undergo esterification reactions. Substitution reactions that replace two chlorine atoms can occur at the same phosphorus atom (giving a *geminal* product) or on different phosphorus atoms. In the latter case, the two substituents can be on the same side of the ring in *cis* positions or on opposite sides of the ring in *trans* positions.

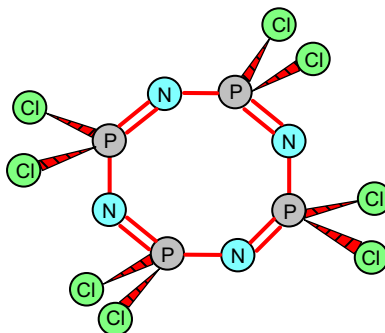


Substitution reactions can produce alkoxide, alkyl, amine, and other derivatives with the following reactions being typical.

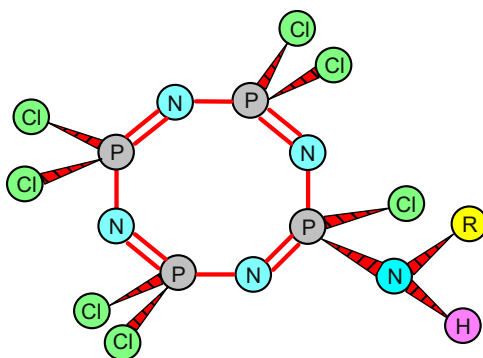




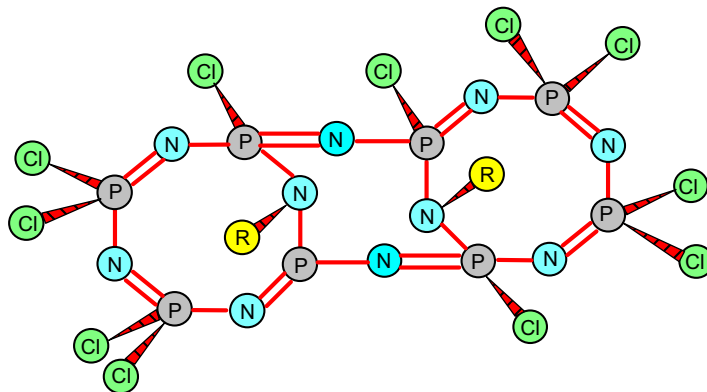
Although the most common phosphazine rings contain three atoms of both phosphorus and nitrogen, other phosphazines are known. For example,  $N_4P_4Cl_8$  has the structure



The chemistry of this compound has been explored by Besli, et al. (2015). For example, this compound reacts with alkyl amines,  $RNH_2$ , to give products having the structure



Under appropriate experimental conditions, sodium hydride, a strong base, will remove the proton from the amine group. This results in the formation of compounds having two of the  $N_4P_4$  ring structures bridged by nitrogen atoms.



Similar reactions are also possible for the cyclotriphosphazene analogs. Although a comprehensive chemistry of phosphazine compounds has been developed, the brief introduction presented here shows only a few of the many types of compounds and reactions in this interesting field.

## 14.8 ACIDS AND THEIR SALTS

Although a number of acids containing arsenic, antimony, and bismuth are known, they are of little importance compared to the acids containing phosphorus. The acid of bismuth,  $H_3BiO_3$  can also be written as  $Bi(OH)_3$  and in keeping with the metallic character of bismuth is it not actually very acidic. Because of their great importance, the discussion that follows is concerned with the acids containing phosphorus.

### 14.8.1 Phosphorous Acid and Phosphites

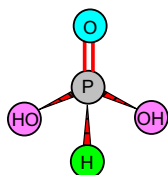
Phosphorous acid,  $\text{H}_3\text{PO}_3$ , is the acid that forms when  $\text{P}_4\text{O}_6$  reacts with water.



For preparing a small amount of the acid in a laboratory, it is more convenient to carefully hydrolyze  $\text{PCl}_3$ .

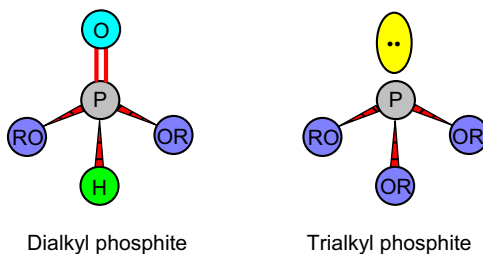


Phosphorous acid is a weak dibasic acid having the molecular structure



The hydrogen atom attached to phosphorus is not acidic so there are only two dissociation constants which have the values  $K_{a1} = 5.1 \times 10^{-2}$  and  $K_{a2} = 1.8 \times 10^{-7}$ . Normal phosphites contain the  $\text{HPO}_3^{2-}$  ion that has an irregular tetrahedral structure. The acid salts contain  $\text{H}_2\text{PO}_3^-$ .

Organic phosphites are relatively important because they have useful properties and function as reactive intermediates for preparing numerous other compounds. The organic phosphites have the general formulas  $(\text{RO})_2\text{P}(\text{O})\text{H}$  and  $(\text{RO})_3\text{P}$  and the structures of these types of compounds are

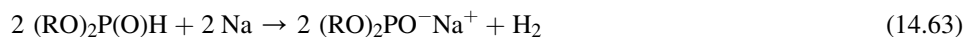


Because the trialkyl phosphates have an unshared pair of electrons on the phosphorus atom, they can function as Lewis bases and thus form complexes with metals. Phosphorus is a soft base so the complexes most often involve a soft metal atom or ion.

The dialkyl phosphites react with chlorine to give a chlorophosphite.



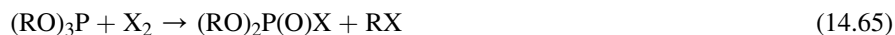
They can also be hydrolyzed to give either a monoalkyl phosphite or free  $\text{H}_3\text{PO}_3$ . A salt that serves as a reactive intermediate is formed by reaction with sodium.



Trialkyl phosphites are prepared by the reaction of  $\text{PCl}_3$  with alcohols,



By controlling the ratio of  $\text{ROH}$  to  $\text{PCl}_3$  the dialkyl phosphites can also be obtained. Compounds of this type are hydrolyzed with the rate of hydrolysis depending on the length of the alkyl chains. Generally, the rate is inversely related to the molecular weight of the phosphite. The phosphites can be halogenated by reactions with  $\text{X}_2$  or  $\text{PX}_3$ .

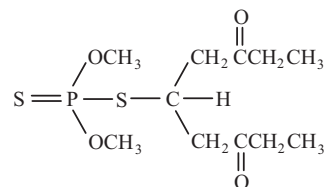


Oxygen, sulfur, or selenium will add to the phosphorus atom to give compounds known as trialkyl phosphates, trialkyl thiophosphates, or trialkyl selenophosphates, respectively, as illustrated by the following equations.

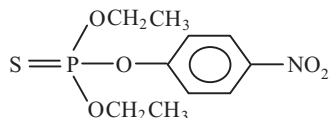


Alkyl phosphites have been used extensively as lubricant additives, corrosion inhibitors, and antioxidants. Their use as intermediates in synthesis provides routes to a wide variety of phosphates, phosphonates, and other organic compounds. Because some of the alkyl phosphites are good solvents for many materials, they have also been useful in solvent extraction processes for separating heavy metals.

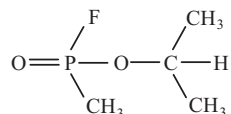
Many organic derivatives of phosphates have been synthesized from the phosphites and are used as insecticides although details of the preparations will not be given here. One such compound is malathion that has the structure



Parathion has the structure



The use of parathion has resulted in many accidental deaths due to its extreme toxicity. In fact, its use in the United States has been prohibited since 1991, but it is still used in other parts of the world. Sarin is a nerve gas that has been produced for military use, and it has the structure



These are only a few of the derivatives of organic phosphates that have been used as toxins. In spite of their toxicity, phosphate derivatives are very useful compounds.

### 14.8.2 Phosphoric Acids and Phosphates

The acids containing phosphorus(V) constitute a complex series that can be considered as resulting from the reaction of  $\text{P}_4\text{O}_{10}$  with varying amounts of water. Some of the acids that are formed by adding  $\text{P}_4\text{O}_{10}$  to water are  $\text{HPO}_3$ ,  $\text{H}_3\text{PO}_4$ , and  $\text{H}_4\text{P}_2\text{O}_7$ . An interesting way to consider the acids is to represent  $\text{P}_4\text{O}_{10}$  in terms of its empirical formula,  $\text{P}_2\text{O}_5$ . Then, the various phosphoric acids represent different ratios of water to  $\text{P}_2\text{O}_5$ . In this way,  $\text{HPO}_3$ , known as metaphosphoric acid, contains the ratio  $\text{H}_2\text{O}/\text{P}_2\text{O}_5$  equal to 1 because the reaction producing  $\text{HPO}_3$  can be shown as



The most “hydrated” compound containing phosphorus is the hypothetical acid  $\text{P}(\text{OH})_5$  that could also be represented by the formula  $\text{H}_5\text{PO}_5$ , and it would be produced by the reaction of 5 moles of water for each mole of  $\text{P}_2\text{O}_5$ .



Thus, we find that the ratio  $\text{H}_2\text{O}/\text{P}_2\text{O}_5$  could theoretically vary from 1 to 5. The  $\text{H}_2\text{O}/\text{P}_2\text{O}_5$  ratio of 3 corresponds to the acid  $\text{H}_3\text{PO}_4$ , and the acids containing condensed phosphates can be considered as being formed from this acid by the loss of

water. Therefore, the practical limit for  $\text{H}_2\text{O}/\text{P}_2\text{O}_5$  is 3. The discussion that follows will show the chemistry of most of these “phosphoric acids” because some of these compounds and their salts are of great economic importance.

Probably the most important phosphorus compound (excluding phosphate fertilizers) is orthophosphoric acid,  $\text{H}_3\text{PO}_4$ . This is the acid intended when the name phosphoric acid is used in most contexts. Approximately 26 billion pounds of this compound are produced annually. The usual commercial form of the acid is a solution containing 85% acid.

Phosphoric acid is prepared by two principal ways that give a product of different purity. Food grade acid is prepared by burning phosphorus and dissolving the product in water.



Phosphoric acid intended for use in manufacturing fertilizers and other chemicals is obtained by treating phosphate rock with sulfuric acid (see Chapter 16).



Liquid  $\text{H}_3\text{PO}_4$  has a high viscosity as a result of it being extensively associated by hydrogen bonding.

Theoretically, it should be possible to add enough  $\text{P}_2\text{O}_5$  (actually it is  $\text{P}_4\text{O}_{10}$ , but the empirical formula is useful here) to the 85% acid to produce 100%  $\text{H}_3\text{PO}_4$ . However, when this is done, it is found that some molecules undergo condensation reactions, one of which can be shown as



Therefore, the solution does not contain 100%  $\text{H}_3\text{PO}_4$  but rather a complex mixture of  $\text{H}_3\text{PO}_4$ ,  $\text{H}_4\text{P}_2\text{O}_7$  (and traces of other acids), and a small amount of  $\text{H}_2\text{O}$ . It has been found that approximately 10% of the total amount of  $\text{P}_2\text{O}_5$  present in what would be expected to be 100%  $\text{H}_3\text{PO}_4$  is actually in the form  $\text{H}_4\text{P}_2\text{O}_7$ . When still more  $\text{P}_2\text{O}_5$  is added, acids such as  $\text{H}_5\text{P}_3\text{O}_{10}$  are formed.

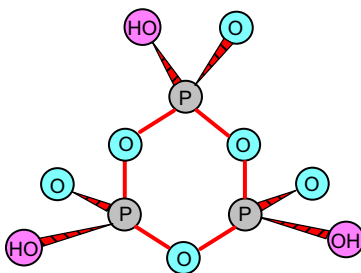
The solution containing 85% phosphoric acid has a density of  $1.686 \text{ g cm}^{-3}$  at  $25^\circ\text{C}$ . As expected from the formula  $(\text{HO})_3\text{PO}$ , phosphoric acid is a weak acid having the dissociation constants as follows:  $K_{a1} = 7.5 \times 10^{-3}$ ;  $K_{a2} = 6.0 \times 10^{-8}$ ;  $K_{a3} = 5 \times 10^{-13}$ . Being a tribasic acid, three series of salts are known that can be written as  $\text{M}_3\text{PO}_4$ ,  $\text{M}_2\text{HPO}_4$ , and  $\text{MH}_2\text{PO}_4$  where M is a univalent ion. Owing to the weak acidity of the acid in the second and third steps of the ionization, salts such as  $\text{Na}_3\text{PO}_4$  form basic solutions resulting from the extensive hydrolysis of the phosphate ion.



Like sulfuric acid, phosphoric acid is not normally an oxidizing agent except at high temperatures.

The list of uses of phosphoric acid and its salts is an impressive one, and only a few of the more general uses will be presented here. The use of phosphoric acid in fertilizers is described in a later section. In dilute solutions, phosphoric acid is nontoxic, and it improves the flavor of carbonated beverages (especially root beer). Because it is much less expensive than citric, tartaric, or lactic acid, it is used in beverages and other food products. The acid is also used in treating metal surfaces as a cleaner, in electroplating, and in phosphate coatings. It is also used as a catalyst in numerous organic syntheses. Of the phosphate salts,  $\text{Na}_2\text{HPO}_4$ ,  $\text{Na}_3\text{PO}_4$ , and  $(\text{NH}_4)_3\text{PO}_4$  are the most important. The first two are used in cleaners, foods, and pigments, whereas the last is an important fertilizer and flameproofing compound.

Metaphosphoric acid,  $\text{HPO}_3$ , has a  $\text{H}_2\text{O}/\text{P}_2\text{O}_5$  ratio of 1. The compound undergoes some association to give a cyclic trimer (sometimes called trimetaphosphoric acid) having the structure

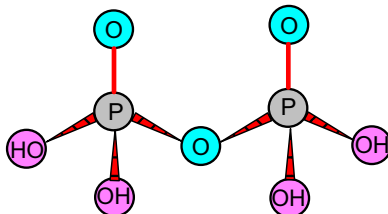


A cyclic tetramer is also known but it is of little practical importance. Metaphosphoric acid is a strong acid as can be seen when the formula is written as  $(\text{HO})\text{PO}_2$  in which, like  $\text{HNO}_3$ , there are two oxygen atoms that do not have hydrogen atoms attached. Some of the metaphosphate salts are used in foods and toiletries.

Pyrophosphoric acid (also known as diphosphoric acid),  $\text{H}_4\text{P}_2\text{O}_7$ , has a mole ratio of  $\text{H}_2\text{O}/\text{P}_2\text{O}_5$  of 2. The equation for its preparation can be written as



The structure of the  $\text{H}_4\text{P}_2\text{O}_7$  molecule is



Formally, the formation of  $\text{H}_4\text{P}_2\text{O}_7$  can be considered as the partial dehydration of  $\text{H}_3\text{PO}_4$  that has the  $\text{H}_2\text{O}/\text{P}_2\text{O}_5$  ratio of 3.



Of course, adding  $\text{P}_2\text{O}_5$  (actually it is  $\text{P}_4\text{O}_{10}$ ) to  $\text{H}_3\text{PO}_4$  accomplishes the same result.

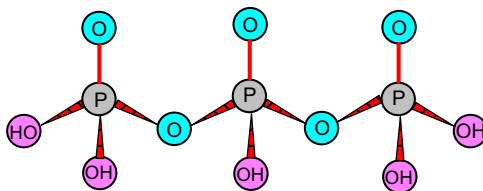


Pyrophosphoric acid (or diphosphoric acid) has four dissociation constants:  $K_{a1} = 1.4 \times 10^{-1}$ ;  $K_{a2} = 1.1 \times 10^{-2}$ ;  $K_{a3} = 2.9 \times 10^{-7}$ ;  $K_{a4} = 4.1 \times 10^{-10}$ . Because the first two dissociation constants are rather close together and much greater than the third and fourth, the first two hydrogen atoms are replaced more easily than the others, and numerous salts of pyrophosphoric acid have the formula  $\text{M}_2\text{H}_2\text{P}_2\text{O}_7$  (where M is a univalent ion). If the solution contains a sufficient amount of base, the other two hydrogens are removed and salts having the formula  $\text{M}_4\text{P}_2\text{O}_7$  result. Salts having the formulas  $\text{MH}_3\text{P}_2\text{O}_7$  and  $\text{M}_3\text{HP}_2\text{O}_7$  are much less numerous.

Other polyphosphoric acids result from the sharing of oxygen atoms on the corner of tetrahedral  $\text{PO}_4$  units. Triphosphoric acid (sometimes called tripolyphosphoric acid),  $\text{H}_5\text{P}_3\text{O}_{10}$ , can be considered as being produced by the reaction

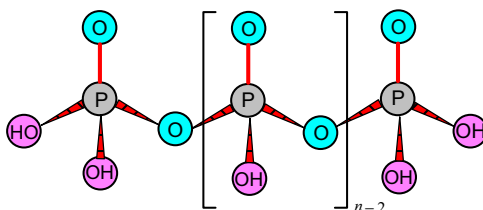


The structure of the acid can be shown as



in which there is a tetrahedral arrangement of four oxygen atoms around each phosphorus atom. It is a strong acid in aqueous solution and the first step of its dissociation is extensive.

Other condensed acids having the general formula  $\text{H}_{n+2}\text{P}_n\text{O}_{3n+1}$  are formed by the elimination of water between the acid containing  $(n - 1)$  phosphorus atoms and  $\text{H}_3\text{PO}_4$ . The general structure of these acids can be represented as



The polyphosphoric acids are not important compounds, but the salts of some of these acids are used extensively. For example, tetrasodium pyrophosphate,  $\text{Na}_4\text{P}_2\text{O}_7$ , is used as a builder in detergents, an emulsifier for making cheese, a dispersant for paint pigments, and in water softening. Sodium dihydrogen pyrophosphate is used as a solid acid that reacts with  $\text{NaHCO}_3$  in baking powder. Tetrapotassium pyrophosphate is used in liquid detergents and shampoos, as a pigment dispersant in paints, and in the manufacture of synthetic rubber. Calcium pyrophosphate is used in tooth paste as a mild abrasive. Tetrasodium pyrophosphate is prepared by the dehydration of solid  $\text{Na}_2\text{HPO}_4$ .



Sodium dihydrogen pyrophosphate is obtained by heating  $\text{NaH}_2\text{PO}_4$  to drive off water.

Of the salts of the polyphosphoric acids,  $\text{Na}_5\text{P}_3\text{O}_{10}$  is by far the most important. An enormous amount of this compound has been used as a builder in sulfonate detergents and as a dispersant. Because the  $\text{P}_3\text{O}_{10}^{5-}$  ion forms stable complexes with  $\text{Ca}^{2+}$  and  $\text{Mg}^{2+}$ , it has been used as a complexing agent in laundry products to prevent these ions from forming precipitates with soap (see [Section 14.1](#)).

## 14.9 PHOSPHORUS IN THE ENVIRONMENT

Phosphorus is an essential constituent in many biological processes. It is found in the nucleic acids RNA and DNA, and phosphate esters form a bridge that binds the double helix in DNA. Phosphorus is also essential in bones and teeth and functions in a buffering system in the blood.

As with nitrogen, the movement of phosphorus through the environment can be described in a cycle known as the *phosphorus cycle*. In the case of phosphorus, this cycle is limited to terrestrial and aquatic environments (as opposed to atmospheric environments) because phosphorus-containing compounds are generally solid at ambient temperature and pressure. In the phosphorus cycle, rocks and minerals release phosphorus into the soil where it is incorporated by plants from which it is subsequently incorporated into animals. When plants are harvested, over time soils become deficient in phosphorus and, as discussed in [Section 14.9.2](#), fertilizers containing phosphorus are added to the soil. Materials that contain an excess of compounds containing phosphorus may result in runoff into lakes and streams where solid phosphates can then precipitate to form a sediment on the bottom.

### 14.9.1 Phosphorus in Soaps and Detergents

In addition to being used in fertilizers, phosphates have also been used in detergents. Hard water contains ions such as calcium, iron, and manganese. With traditional forms of soap (sodium or potassium salts of long chain organic acids),  $\text{Ca}^{2+}$ ,  $\text{Fe}^{2+}$ , and  $\text{Mn}^{2+}$  ions bind to the organic portion of the soap molecule to form a precipitate known as “soap scum.” Until recently, phosphate-containing compounds were added to detergents as *builders* that form complexes with metal ions found in hard water and thus prevent the formation of soap scum. Generally, sodium triphosphate,  $\text{Na}_5\text{P}_3\text{O}_{10}$ , was added to dry detergents and sodium or potassium phosphates were added to liquid detergents. By the mid-1980s, approximately 2 million tons of phosphorus was used annually in detergents in the United States.

As with excess nitrogen in saltwater estuaries, excess phosphorus can lead to algae blooms and eutrophication in freshwater sources. In 2000, the use of phosphates in laundry detergents was discontinued in the United States, and in 2010, phosphates were eliminated from dishwasher detergents.

### 14.9.2 Fertilizer Production

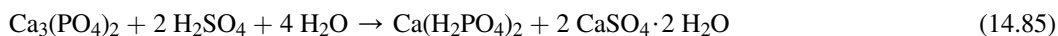
With a world population that has reached six billion, the production of food represents a monumental problem. Without the use of effective fertilizers, it would be impossible to meet this demand. For the most part, the fertilizers that are needed in such huge quantities are inorganic materials.

Enormous quantities of phosphorus compounds are used in the production of fertilizers. Calcium phosphate is found in many regions of the world, but its direct use as a fertilizer is not very effective because of its low solubility. As was mentioned in Chapter 1, sulfuric acid plays an important role in fertilizer production and approximately 65% of the sulfuric

acid manufactured (over 80 billion pounds annually) is used for this purpose. Sulfuric acid is produced by the following reactions.

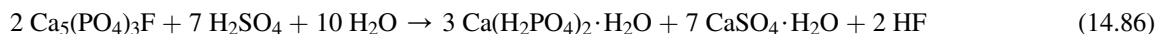


When pulverized calcium phosphate is treated with sulfuric acid (the least expensive strong acid), the reaction is



Approximately 100 billion pounds of phosphate rock are processed annually, primarily to produce fertilizers. This reaction occurs because the  $\text{PO}_4^{3-}$  is the conjugate base of a weak acid and it is easily protonated. The mixture of calcium dihydrogen phosphate and calcium sulfate (*gypsum*) is called *superphosphate of lime*, and it contains a higher percentage of phosphorus than does calcium phosphate,  $\text{Ca}_3(\text{PO}_4)_2$ . Moreover, it contains the phosphate in a soluble form,  $\text{Ca}(\text{H}_2\text{PO}_4)_2$ . Note that the latter compound contains a +2 cation and two -1 anions, whereas the former contains +2 cations and -3 anions. The lattice composed of the lower charged ions is more readily separated by a solvent (water in this case) than is one that contains ions having higher charges on the ions.

Some fluoroapatite,  $\text{Ca}_5(\text{PO}_4)_3\text{F}$ , is often found along with  $\text{Ca}_3(\text{PO}_4)_2$ . Fluoroapatite reacts with sulfuric acid according to the following equation.



Therefore, in addition to producing fertilizer, this reaction is also a source of hydrogen fluoride.

Phosphoric acid is prepared by the reaction of calcium phosphate with sulfuric acid.



Other important fertilizers are obtained by the reactions of phosphoric acid with calcium phosphate and fluoroapatite. These processes can be represented by the following equations.

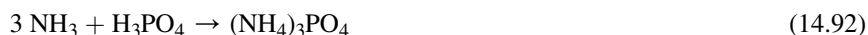


The compound  $\text{Ca}(\text{H}_2\text{PO}_4)_2$  contains a higher percentage of phosphorus than does the mixture with calcium sulfate known as *superphosphate of lime*. Accordingly, it is sometimes called *triple superphosphate*.

In addition to fertilizers that provide phosphorus, sources of nitrogen compounds are also needed. One material used in large quantities is ammonium nitrate. In Chapter 13, the commercial preparations of nitric acid and ammonia were described. These compounds react directly to produce ammonium nitrate that is not only used as a fertilizer but also as an explosive (see Chapter 13). About 14 billion pounds of ammonium nitrate are produced annually primarily by the reaction



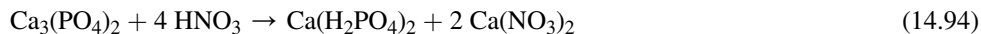
Other ammonium salts are used as fertilizers, and the first step in their production is the production of ammonia by the Haber process. Some of the ammonium salts that are effective for use as fertilizers are ammonium sulfate and ammonium phosphate that are prepared by the following reactions.



Urea is still used in large quantities as a fertilizer (about 13 billion pounds annually) and it is prepared by the reaction of ammonia with carbon dioxide.



Finally, the reaction of calcium phosphate with nitric acid produces two calcium salts that both have value as nutrients in agriculture.



The chemistry described in this section is carried out on a large scale and it is impossible to over emphasize its importance to our way of life. If the projections that indicate that the world's current population of 6 billion will grow to 12 billion by the year 2030 are correct, it is clear that this type of chemistry will become even more important.

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## PROBLEMS

- Write balanced equations for the reactions of  $\text{PCl}_3$  with each of the following compounds.
  - $\text{H}_2\text{O}$
  - $\text{H}_2\text{O}_2$
  - $\text{LiC}_4\text{H}_9$
  - $\text{CH}_3\text{OH}$
- Complete and balance the following.
  - $\text{As}_2\text{O}_3 + \text{HCl} \rightarrow$
  - $\text{As}_2\text{O}_3 + \text{Zn} + \text{HCl} \rightarrow$
  - $\text{Sb}_2\text{S}_3 + \text{O}_2 \rightarrow$
  - $\text{SbCl}_3 + \text{H}_2\text{O} \rightarrow$
  - $\text{Sb}_2\text{O}_3 + \text{C} \xrightarrow{\text{heat}}$
- Explain the nature of the driving force for the reaction
 
$$\text{PCl}_3 + \text{AsF}_3 \rightarrow \text{PF}_3 + \text{AsCl}_3$$
- Draw structures for the following species.
  - Pyrophosphoric acid
  - $\text{P}_4\text{O}_{10}$
  - Trimetaphosphoric acid
  - $\text{P}_2\text{O}_7^{4-}$



5. Draw structures for the following species.
  - (a) Phosphorous acid
  - (b) Solid  $\text{PCl}_5$
  - (c) Hypophosphorous acid
  - (d) Phosphoryl chloride
6. Complete and balance the following.
  - (a)  $\text{NaCl} + \text{SbCl}_3 \xrightarrow{\text{heat}}$
  - (b)  $\text{Bi}_2\text{S}_3 + \text{O}_2 \xrightarrow{\text{heat}}$
  - (c)  $\text{Na}_3\text{Sb} + \text{HCl} \rightarrow$
  - (d)  $\text{BiBr}_3 + \text{H}_2\text{O} \rightarrow$
  - (e)  $\text{Bi}_2\text{O}_3 + \text{C} \xrightarrow{\text{heat}}$
7. Phosphoric acid is prepared in two ways commercially. Both methods start with calcium phosphate as the raw material.
  - (a) Write a complete set of equations to show the preparation of impure phosphoric acid.
  - (b) Write a complete set of equations to show the preparation of pure phosphoric acid.
8. Although phosphorous acid has the formula  $\text{H}_3\text{PO}_3$ , the titration with sodium hydroxide gives  $\text{Na}_2\text{HPO}_3$ . Explain why.
9. Draw structures for two other species that are isoelectronic with  $\text{P}_3\text{O}_9^{3-}$ .
10. Write balanced equations for each of the following processes.
  - (a) The preparation of  $(\text{CH}_3\text{O})_3\text{PO}$
  - (b) The preparation of  $\text{POF}_3$  (starting with  $\text{P}_4$ )
  - (c) The commercial preparation of phosphorus
  - (d) The reaction of  $\text{P}_4$  with  $\text{NaOH}$  and water
  - (e) The preparation of superphosphate fertilizer
  - (f) The preparation of phosphine (starting with phosphorus)
11. Write balanced equations for each of the following processes.
  - (a) The preparation of triethylphosphate
  - (b) The reaction of  $\text{PCl}_5$  with  $\text{NH}_4\text{Cl}$
  - (c) The preparation of arsine
  - (d) The reaction of  $\text{P}_4\text{O}_{10}$  with  $i\text{-C}_3\text{H}_7\text{OH}$
12. Starting with elemental phosphorus, show a series of equations to synthesize the following.
  - (a)  $\text{P}(\text{OCH}_3)_3$
  - (b)  $\text{OP}(\text{OC}_2\text{H}_5)_3$
  - (c)  $\text{SP}(\text{OC}_2\text{H}_5)_3$
13. Complete and balance the following.
  - (a)  $\text{SbCl}_3 + \text{LiC}_4\text{H}_9 \rightarrow$
  - (b)  $\text{PCl}_5 + \text{P}_4\text{O}_{10} \rightarrow$
  - (c)  $\text{OPBr}_3 + \text{C}_2\text{H}_5\text{OH} \rightarrow$
  - (d)  $(\text{NPCl}_2)_3 + \text{LiCH}_3 \rightarrow$
  - (e)  $\text{P}(\text{OC}_2\text{H}_5)_3 + \text{S} \rightarrow$
14. Write the balanced equations for the following processes.
  - (a) The reaction of  $\text{As}_4\text{O}_6$  with  $\text{HF}$
  - (b) The preparation of  $\text{H}_4\text{P}_2\text{O}_7$  starting with  $\text{P}_4$ , air, and water
  - (c) The reaction of  $\text{SO}_2$  with  $\text{PCl}_5$
  - (d) The commercial preparation of urea
  - (e) The dehydration of disodium hydrogen phosphate
15. In the reaction of  $\text{PF}_2\text{I}$  with mercury, explain why it is iodine rather than fluorine that is removed.
16. There are two possible structures for a molecule that contains one phosphorus atom and three cyanide groups. Which structure is most stable? Why would the other possible structure be less stable?

## Chapter 15

# Oxygen

Oxygen and its compounds are widespread and common. An enormous number of oxygen compounds are found in nature, and the earth's crust contains large amounts of oxygen-containing compounds. For example, the atmosphere is 21% oxygen by volume or 23.3% by weight of this element. Water, which covers approximately two-thirds of the earth's surface, contains 89% oxygen. Oxygen itself is necessary for the life processes involving all known animal forms of life. Many of the solid rocks and minerals are complex oxygen-containing compounds in the form of silicates, phosphates, and carbonates.

Oxygen and its compounds are also of vital importance in commerce and industry. Many of the compounds which are used in the largest quantities contain oxygen. These chemicals are not laboratory curiosities, but rather they are important raw materials in many of the manufacturing processes which lead to objects that we now regard as necessary. In view of this, it is easy to see that some knowledge of the chemistry of oxygen is necessary for an understanding of our surroundings.

It is readily apparent that oxygen and its compounds are used in quantities that are almost beyond comprehension. Sulfuric acid is by far the leading chemical in terms of production. This compound is used in numerous manufacturing processes both in heavy industry and in specialty preparations and consists of over 65% oxygen. Sulfuric acid is so widely used that it has been stated that the production and use of the chemical provides a "barometer" for gauging the status of the economy. In the sections that follow, we will explore some of the chemistry of this most important element, oxygen.

### 15.1 ELEMENTAL OXYGEN, O<sub>2</sub>

Oxygen is a colorless, odorless, tasteless gas that is slightly soluble in water. The solubility increases with increasing pressure and decreases with increasing temperature. This solubility of oxygen provides the basis for the existence of aquatic animal life.

Naturally occurring oxygen consists of three isotopes. These have the mass numbers of 16, 17, and 18, and they are present in the percentages of 99.759, 0.037, and 0.204, respectively. Prior to 1961, <sup>16</sup>O was the standard of atomic mass and it was taken as exactly 16.0000 amu by physicists. However, chemists usually considered the natural mixture of isotopes as being 16.0000 amu, but because of the presence of small amounts of <sup>17</sup>O and <sup>18</sup>O, the average is actually slightly greater than 16.0000 amu. Thus, there were two different atomic mass scales. As we described in Chapter 11, this is no longer the case because <sup>12</sup>C is now the basis for the atomic mass scale, and its mass is assigned as 12.0000 amu. On this scale, <sup>16</sup>O has a mass of 15.994915 amu.

The oxygen molecule is diatomic, and the molecular orbital (MO) diagram for the molecule is shown in Figure 15.1.

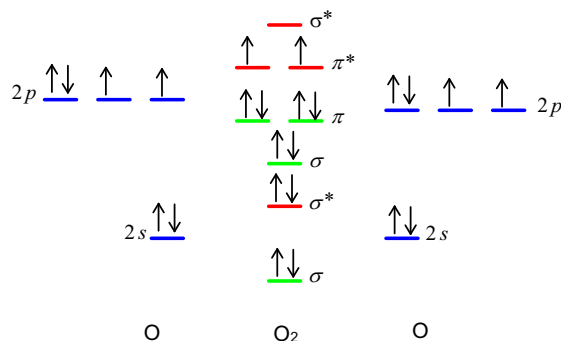


FIGURE 15.1 Molecular orbital diagram for O<sub>2</sub>.

**TABLE 15.1** Characteristics of Dioxygen Species

Property	O <sub>2</sub> <sup>+</sup>	O <sub>2</sub>	O <sub>2</sub> <sup>-</sup>	O <sub>2</sub> <sup>2-</sup>
Bond length, pm	112	121	128	149
Bond energy, kJ mol <sup>-1</sup>	623	494	393	142
Force constant, mdyne Å <sup>-1</sup>	16.0	11.4	5.6	4.0

From the MO diagram shown, it is readily apparent that O<sub>2</sub> is paramagnetic, having two unpaired electrons per molecule. These unpaired electrons reside in antibonding orbitals, and they are responsible for O<sub>2</sub> reacting as a free radical in many instances. Removal of one or both of these electrons from O<sub>2</sub> results in O<sub>2</sub><sup>+</sup> or O<sub>2</sub><sup>2+</sup>, both of which have shorter bond distances than O<sub>2</sub>. Adding one or two electrons in the π\* orbitals results in the formation of O<sub>2</sub><sup>-</sup> and O<sub>2</sub><sup>2-</sup>, respectively. All of these species are known and the trends in their bond lengths and bond strengths are shown by the data in Table 15.1.

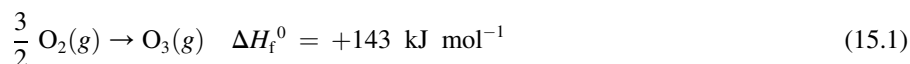
Elementary chemistry books sometime show a valence bond structure for the O<sub>2</sub> molecule as



This structure is, of course, incorrect because it does not account for the fact that the O<sub>2</sub> molecule is paramagnetic, having two unpaired electrons per molecule. An O–O single bond has an energy of about 142 kJ mol<sup>-1</sup> whereas the O=O bond energy is about 494 kJ mol<sup>-1</sup>. The actual bond energy in the O<sub>2</sub> molecule is very close to the double bond energy, which is consistent with the bond order of 2 found from the MO diagram above. Therefore, there is no satisfactory way to draw a valence bond diagram for the O<sub>2</sub> molecule that is consistent with its paramagnetism and still obeys the octet rule. The fact that O<sub>2</sub> is paramagnetic is easily demonstrated because liquid oxygen can be suspended between the poles of a powerful magnet. In the liquid state, oxygen has a light blue color owing to absorption of light in the visible region due to electron transitions between the π\* and σ\* antibonding orbitals in the molecule.

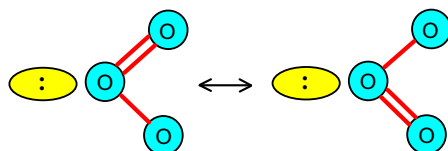
## 15.2 OZONE, O<sub>3</sub>

In addition to the usual form of oxygen, there exists a second allotrope, ozone, O<sub>3</sub>. This is a high energy form of the element having a heat of formation of +143 kJ mol<sup>-1</sup>.



In 1785, a Dutch scientist, Van Marum noted an unusual odor in oxygen that had an electric spark passed through it. He also noted that the gas reacted easily with mercury. Schonbein showed in 1840 that the gas liberates iodine from potassium iodide solutions. He gave it the name ozone, the origin of which is the Greek word meaning “to smell.” Some years later, it was shown that the gas was produced from the oxygen itself.

Liquid ozone is a blue, diamagnetic substance which is explosive when mixed at high concentrations with oxygen. The explosive character of O<sub>3</sub>/O<sub>2</sub> mixtures increases at high pressure so that the mixture is maintained at a temperature lower than -20 °C. Ozone is black as a solid, and it has a melting point of -193 °C and a boiling point of -112 °C. Ozone is somewhat more soluble in water than is O<sub>2</sub> with a given volume of water dissolving about half that volume of O<sub>3</sub> at 0 °C whereas only about 0.049 volumes of O<sub>2</sub> will dissolve. The structure of the molecule is angular (C<sub>2v</sub>) with two principle resonance structures that can be shown as follows.



These structures are in accord with O<sub>3</sub> being an 18-electron triatomic molecule. The bond angle is 117° and the bond lengths are 128 pm.

Small concentrations of ozone are found in the upper atmosphere where it is produced by sunlight in the ultraviolet region of the spectrum. Ozone absorbs ultraviolet light in the 200–360 nm (2000–3600 Å) range which provides protection from intense UV radiation at the earth's surface. Chlorofluorocarbons, used as a propellant in aerosol cans and as a refrigerating gas, are thought to cause its destruction by a complicated series of reactions. This concern for the ozone layer has led to the banning of chlorofluorocarbons for many uses. Of course, an increase in UV radiation at the earth's surface would lead to an increase in skin cancer, among other things.

Ozone can be produced by electric discharge through oxygen as illustrated by the equation,



which results in a few percent conversion. Its decomposition is catalyzed by  $\text{Na}_2\text{O}$ ,  $\text{K}_2\text{O}$ ,  $\text{MgO}$ ,  $\text{Al}_2\text{O}_3$ ,  $\text{Cl}_2$ , and other substances.

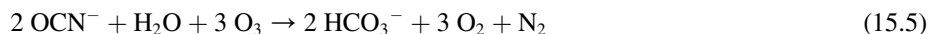
Ozone will add an electron to produce an *ozonide*,  $\text{O}_3^-$ , which is isoelectronic with  $\text{ClO}_2$  (it contains a total of 19 electrons). Some ozonides are stable compounds. For example  $\text{KO}_3$  has a heat of formation of  $-260 \text{ kJ mol}^{-1}$ . This red, paramagnetic compound is stable to  $60^\circ\text{C}$  and is a very strong oxidizing agent. It reacts with water to liberate oxygen.



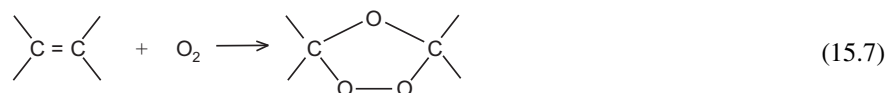
However, the compound undergoes slow decomposition to form the superoxide,



Ozone is a useful oxidizing agent. It reacts to produce  $\text{O}_2$  as the reduction product so that it does not introduce undesirable contaminants. It is comparable to fluorine and atomic oxygen in strength as an oxidizing agent. A minor use of ozone is in the oxidation of metals to highest oxidation states where certain metals are more easily separated. The major use of ozone is in water purification. It is a "clean" oxidant and a potent germicide. It will destroy a variety of species such as cyanide, cyanate, etc., through the following type of reactions:



Another of the important uses of ozone is in reactions with many organic compounds, especially those containing double bonds,



to produce ozonides. Compounds of this type are precursors to products containing other functional groups.

### 15.3 PREPARATION OF OXYGEN

Because the atmosphere is about 21% oxygen, this is the usual source of oxygen on a large scale. Liquid air is distilled and, because the boiling point of oxygen is  $-183^\circ\text{C}$ , nitrogen, which has a boiling point of  $-196^\circ\text{C}$ , boils off first. Smaller amounts of oxygen are obtained from the electrolysis of water, which is carried out to produce high purity hydrogen.

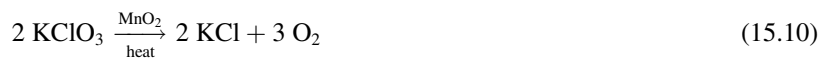
On a laboratory scale,  $\text{O}_2$  can be prepared by decomposition of certain oxygen-containing compounds. Heavy metal oxides, such as  $\text{HgO}$ , are not thermally stable and are fairly easily decomposed.



Peroxides can also be decomposed.



Finally, the usual general chemistry experiment for the preparation of oxygen involves the decomposition of  $\text{KClO}_3$ .



This reaction is very complex with part of the  $\text{MnO}_2$  “catalyst” being converted to the permanganate,  $\text{KMnO}_4$ , which then decomposes.



Eventually, the manganate also decomposes so that  $\text{MnO}_2$  is regenerated in the end. The process is considerably more complicated than the straightforward decomposition of  $\text{KClO}_3$ .

Electrolysis of aqueous solutions of many compounds leads to the preparation of oxygen because  $\text{OH}^-$  is more easily oxidized than are many other anions. For example,



The oxygen is, of course, liberated at the anode as a result of the process being an oxidation reaction.

## 15.4 BINARY COMPOUNDS OF OXYGEN

### 15.4.1 Ionic (Metallic) Oxides

Although adding one electron to an oxygen atom to produce  $\text{O}^-$  liberates about  $142 \text{ kJ mol}^{-1}$ , adding a second electron absorbs enough energy so that the process



absorbs  $703 \text{ kJ mol}^{-1}$ . However, many ionic oxides containing  $\text{O}^{2-}$  exist, and the heats of formation are negative owing to the large amount of energy released when the lattice forms. Oxygen forms one or more binary compounds with almost all elements in the periodic chart. Those with metals such as Ag, Hg, etc., are not particularly stable due to the mismatch in the size of the cation and anion and the fact that these metals are easily reduced. Oxygen compounds with small, highly charged cations (e.g.,  $\text{Mg}^{2+}$ ,  $\text{Fe}^{3+}$ ,  $\text{Cr}^{3+}$ ) are very stable as characterized by their high melting points and decomposition temperatures. Metal oxides are particularly important if for no other reason than because there are over 100 known elements about three-fourths of which are metals.

Reactions of the Group IA metals with oxygen are particularly interesting as shown by the following reactions:



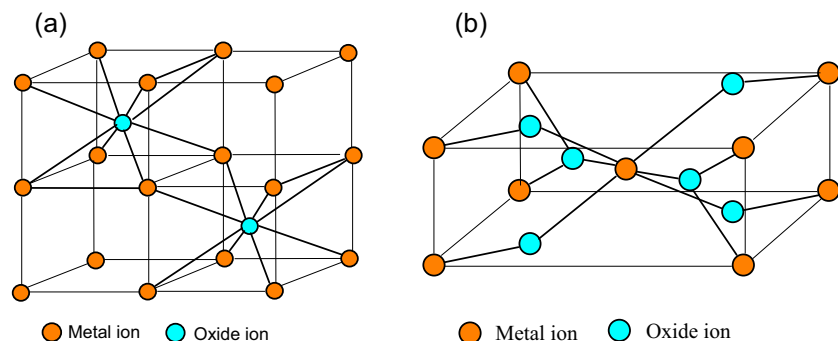
Reactions of Rb and Cs with oxygen also give superoxides. Group II metals follow a similar pattern. Be, Mg, Ca, and Sr give normal oxides, and Ba gives predominantly the peroxide. Apparently, Ra gives either a peroxide or some superoxide depending on the reaction conditions.

Many metals are found in nature as the oxides (e.g.,  $\text{Fe}_2\text{O}_3$ ). Because it has an extremely high heat of formation,  $\text{Al}_2\text{O}_3$  is extremely stable. Therefore, even though  $\text{Fe}_2\text{O}_3$  is itself very stable, the reaction



is strongly exothermic. It produces enough heat that the iron is produced in the molten state. This reaction is known as the *thermite reaction*. Replacement of the  $\text{Fe}^{3+}$  ion by the  $\text{Al}^{3+}$  ion (a smaller, harder, less polarizable ion) would be expected to be favorable on the basis of the hard–soft interaction principle (see Chapter 6).

Metal oxides that have the general formula  $\text{M}_2\text{O}$  most often have the antifluorite structure shown in [Figure 15.2\(a\)](#) whereas those having the formula  $\text{MO}_2$  often have a structure of the rutile type shown in [Figure 15.2\(b\)](#). Oxides having the general formula  $\text{MO}$ , such as  $\text{FeO}$ , most often have a structure of the NaCl type. Other metal oxides, such as  $\text{Fe}_2\text{O}_3$ , have multiple crystal structures that can be often be interchanged under some conditions.



**FIGURE 15.2** The antifluorite crystal structure of  $\text{Li}_2\text{O}$ ,  $\text{Na}_2\text{O}$ ,  $\text{K}_2\text{O}$ , and  $\text{Rb}_2\text{O}$  (a) and the rutile structure (b) of compounds having the formula  $\text{MO}_2$ .

Binary oxides are important in the study of the chemistry of acids and bases. For example, the ionic oxides are the anhydrides of strong bases as illustrated by the equations



However, calcium hydroxide, like the hydroxides of most heavy metals and transition metals, is not very soluble in water. Calcium oxide is produced by heating calcium carbonate,



Calcium oxide used in industrial applications is known as *lime*, and calcium hydroxide is known as *hydrated lime* or *slaked lime*. The oxide and the hydroxide are extremely important bases which are used on an enormous scale because they are less expensive than  $\text{NaOH}$ . In the chemical industry, one would rarely use  $\text{NaOH}$  if  $\text{CaO}$  or  $\text{Ca}(\text{OH})_2$  could be used.

The oxide ion is such a strong base that it reacts with water,



so that ionic oxides are usually not products of reactions carried out in aqueous solutions.

### 15.4.2 Covalent Oxides

Oxygen forms covalent oxides with most of the nonmetallic elements and with metallic elements in high oxidation states. The permanganate ion,  $\text{MnO}_4^-$ , is tetrahedral, and it would certainly be considered covalent in view of the fact that  $\text{Mn}$  has a +7 oxidation state. Combustion in air (with an excess of oxygen, the oxidizing agent) generally produces a nonmetal oxide in which the nonmetal is in the highest oxidation state possible. Although this is true for elements such as phosphorus and carbon, it is not true for sulfur which produces  $\text{SO}_2$ , not  $\text{SO}_3$ .

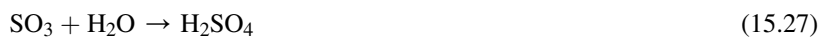


Combustion of a nonmetal in a deficiency of oxygen (the reducing agent is in excess) produces an oxide in which the nonmetal is in some lower oxidation state.



The structures of the covalent oxides are of considerable interest, and they will be discussed as part of the chemistry of the central atom.

As the following equations show, the covalent oxides are the anhydrides of acids.



Because the oxides of nonmetals are usually acidic whereas those of metals are basic, reactions between the oxides themselves may take place to produce salts directly.



### 15.4.3 Amphoteric Oxides

The oxides of some elements are amphoteric, meaning they are neither characteristically acidic nor basic. Particularly interesting are the oxides of Zn and Al. For example,

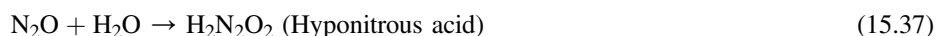
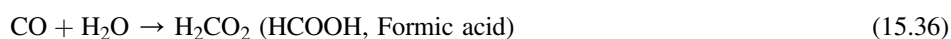


In the first of these reactions, ZnO behaves as a base, and in the second it behaves as an acid. This behavior can also be shown for the hydroxides.



The oxides of the second long period of elements can be arranged in a continuum to illustrate the transition in acidic and basic properties as is shown in Figure 15.3. It is also interesting to consider the effects that arise as the oxides in a given group are examined with respect to acidic and basic character. If one considers the series  $\text{CO}_2$ ,  $\text{SiO}_2$ ,  $\text{GeO}_2$ , and  $\text{PbO}_2$ , the first is a weakly acidic oxide whereas the last is a weakly basic oxide.

A few oxides are not anhydrides of either acids or bases. For example, CO and  $\text{N}_2\text{O}$  do not react directly with water although they are formally the anhydrides of acids. Therefore, we can write the equations



but these reactions cannot be carried out in this way.

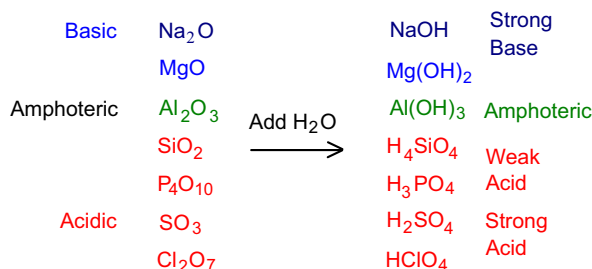


FIGURE 15.3 Oxides of elements classified in terms of their reactions with water.

### 15.4.4 Peroxides and Superoxides

We have already seen that the alkali metals do not all give usual or normal oxides when they react with excess oxygen. In a deficiency of oxygen, however, the oxides (containing  $O^{2-}$ ) are obtained. Ionic peroxides contain the  $O_2^{2-}$  ion (oxygen is in the  $-1$  oxidation state). The peroxides of sodium, barium, and radium can be formed by direct reaction of those elements with oxygen, but lithium peroxide is synthesized by reaction of lithium hydroxide with hydrogen peroxide according to a reaction represented by the equation



When heated, peroxides decompose to form normal oxides and oxygen gas. For example,



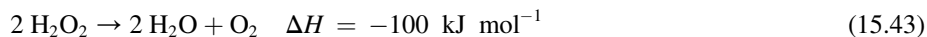
The  $O_2^{2-}$  ion is a base so peroxides accept protons from water to give  $\text{H}_2\text{O}_2$ .



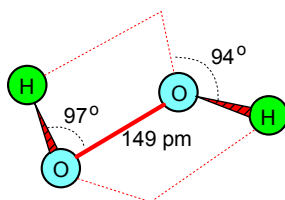
Peroxides are also very strong oxidizing agents. Moreover, because the  $O_2^{2-}$  ion is a base, peroxides have the ability to react with acidic oxides. For example, CO and  $\text{CO}_2$  react with  $\text{Na}_2\text{O}_2$  as represented by the following equations:



Hydrogen peroxide is relatively unstable and decomposes to give water and oxygen as shown in the following equation:



Hydrogen peroxide cannot be concentrated by distillation from aqueous solutions above about 30% because at that concentration decomposition takes place as rapidly as distillation. The decomposition is catalyzed by transition metal ions so that containers for concentrated solutions of  $\text{H}_2\text{O}_2$  must be very clean. Very concentrated  $\text{H}_2\text{O}_2$  (>90%) is a strong oxidant, and it has been used as the oxidizing agent in a high energy rocket fuel. The usual 3%  $\text{H}_2\text{O}_2$  marketed in retail stores is used as a bleach and disinfectant. The structure of the  $\text{H}_2\text{O}_2$  molecule can be shown as follows.



Pure  $\text{H}_2\text{O}_2$  is a colorless liquid having a melting point of  $-0.43^\circ\text{C}$  and a boiling point of  $150.2^\circ\text{C}$ . It is prepared by electrolyzing cold concentrated  $\text{H}_2\text{SO}_4$  to convert it to peroxydisulfuric acid,  $\text{H}_2\text{S}_2\text{O}_8$ . Hydrolysis of  $\text{H}_2\text{S}_2\text{O}_8$  gives  $\text{H}_2\text{O}_2$ ,



which is then separated and concentrated by distillation. Organic peroxides,  $\text{R-O-O-R}$ , are produced by the reaction of oxygen with ethers,



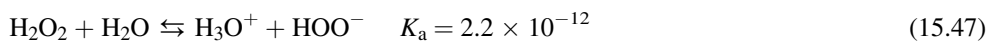
Many of these compounds that contain peroxide linkages are dangerous explosives, but they can be hydrolyzed to produce hydrogen peroxide. Finally,  $\text{H}_2\text{O}_2$  can be prepared by the action of acids on  $\text{BaO}_2$ .



Hydrogen peroxide is a useful and versatile oxidizing agent, but it must be handled carefully when the concentration is 30% or above. The fact that its decomposition is catalyzed by iron and manganese ions makes it essential that it be kept very pure and cautious use and storage be observed.



Hydrogen peroxide is a very weak acid which gives rise to the reaction



As a result, the only way to obtain substances that contain the  $\text{H}_3\text{O}_2^+$  ion is by the use of extremely strong acids such as  $\text{HSbF}_6$ .



The elements sodium, potassium, rubidium, and cesium react with oxygen to form what are known as superoxides, which contain the  $\text{O}_2^-$  ion. Although the superoxides are stable under dry conditions, they will react with water to form hydroxides and oxygen gas.



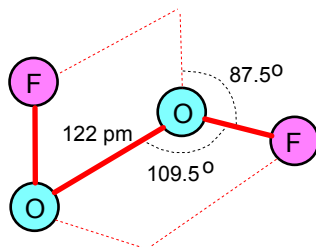
This reaction has been used to generate pure oxygen in spacecraft and in submarines. For superoxides, the oxygen is in the  $-1/2$  oxidation state.

## 15.5 POSITIVE OXYGEN

Because oxygen is the second most electronegative element, it can assume positive oxidation states only in combination with the element having the highest electronegativity, fluorine. Of the oxygen compounds with fluorine, the most thoroughly studied is  $\text{OF}_2$ , oxygen difluoride, which can be prepared by the following reaction:



The structure of  $\text{OF}_2$  is angular ( $C_{2v}$  symmetry) with a bond angle of  $103^\circ$ . It is a pale yellow compound having a boiling point of  $-145^\circ\text{C}$ . Other binary compounds of oxygen and fluorine include  $\text{O}_2\text{F}_2$  and  $\text{O}_3\text{F}_2$ . Dioxygen difluoride,  $\text{O}_2\text{F}_2$ , is analogous to hydrogen peroxide, and it has a structure similar to that of the hydrogen peroxide molecule.



The ionization potential of the oxygen atom is  $13.6 \text{ eV}$  ( $1312 \text{ kJ mol}^{-1}$ ). Therefore, although oxygen *atoms* are formally positive in the oxygen fluorides described above, they are not present as positive ions of oxygen. However, we have seen that some compounds contain the  $\text{NO}^+$  ion. The ionization energy for the  $\text{NO}$  molecule is only  $9.23 \text{ eV}$  ( $891 \text{ kJ mol}^{-1}$ ), with the electron being removed from an antibonding ( $\pi^*$ ) orbital. Molecular oxygen has an ionization potential of  $12.08 \text{ eV}$  ( $1166 \text{ kJ mol}^{-1}$ ). Accordingly, it is not unreasonable to expect that a very potent oxidizing agent could remove an electron from an  $\text{O}_2$  molecule to produce the  $\text{O}_2^+$  cation. The interaction of  $\text{PtF}_6$  with molecular oxygen provides such a reaction.



The  $\text{O}_2\text{PtF}_6$  contains the  $\text{O}_2^+$  cation and is isomorphous with  $\text{KPtF}_6$ . The  $\text{O}_2^+$  cation is called the dioxygenyl cation, and it can also be prepared by the reaction



Molecular parameters for the  $\text{O}_2^+$  cation can be found in [Table 15.1](#).

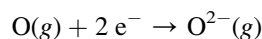
The ionization potential of xenon,  $\text{Xe}$ , is  $12.127 \text{ eV}$  ( $1170 \text{ kJ mol}^{-1}$ ). The fact that this ionization potential is almost exactly the same as that of  $\text{O}_2$  suggests that  $\text{Xe}$  might react similarly to produce a compound containing a noble gas. In the early 1960s, Neil Bartlett carried out such a reaction and obtained a product containing  $\text{Xe}$ , as will be discussed in Chapter 18.

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## PROBLEMS

- The process of adding two electrons to an oxygen atom,



absorbs 703 kJ mol<sup>-1</sup>. Why, therefore, do so many ionic oxides exist?

- Draw MO energy level diagrams for O<sub>2</sub><sup>-</sup> and O<sub>2</sub><sup>2-</sup> and explain how they reflect the chemistry of these species.
- Suggest an explanation for why some metals are found in nature as the oxides whereas others (such as lead and mercury) are found as sulfides.
- Complete and balance the following.
  - MgO + SiO<sub>2</sub>  $\xrightarrow{\text{heat}}$
  - Na<sub>2</sub>S<sub>2</sub>O<sub>8</sub>  $\xrightarrow{\text{heat}}$
  - KO<sub>3</sub> + H<sub>2</sub>O →
  - Cr<sub>2</sub>O<sub>3</sub> + Al  $\xrightarrow{\text{heat}}$
  - ZnO + OH<sup>-</sup> + H<sub>2</sub>O →
- Explain why, of the alkali metals, only lithium forms a normal oxide.
- Describe the commercial process for the production of H<sub>2</sub>O<sub>2</sub>.
- Describe the SO molecule by drawing a MO energy level diagram. Comment on the stability and nature of this molecule.
- The melting points for KO<sub>2</sub>, RbO<sub>2</sub>, and CsO<sub>2</sub> are 380, 412, and 432 °C, respectively. Explain this trend in melting points.
- Explain why ozone is useful in water purification.
- Draw the MO energy level diagram for O<sub>2</sub> and describe what happens to the properties of the molecule when an electron is removed.
- Chemically, oxygen is a very reactive gas while nitrogen is comparatively inert. Explain this difference.
- Molecular oxygen forms numerous complexes with transition metals. Speculate on the bonding modes that could occur without rupture of the O—O bond.
- The species O<sub>2</sub><sup>+</sup>, O<sub>2</sub><sup>-</sup>, and O<sub>2</sub><sup>2-</sup> have stretching frequencies in the infrared region at 1145, 842, and 1858 cm<sup>-1</sup> but not in that order. Match the stretching frequencies to the appropriate species and explain your line of reasoning. Based on your assignments, estimate the position of the O—O stretching frequency in H<sub>2</sub>O<sub>2</sub>.
- The bond angles in H<sub>2</sub>O, HOF, and OF<sub>2</sub> are 104.5, 97, and 103°, respectively. In terms of the orbitals used and atomic properties, explain this trend.
- Mixtures of H<sub>2</sub>S and OF<sub>2</sub> react explosively. Write an equation for the reaction and explain why it is so energetically favorable.
- Superoxides are generally yellow in color. Draw an orbital energy diagram and suggest a possible origin for the absorption that could give rise to the color.

# Sulfur, Selenium, and Tellurium

In this chapter, the chemistry of sulfur, selenium, and tellurium will be described. Polonium will be mentioned only briefly in keeping with the fact that all of the isotopes of the element are radioactive. As a result, the chemistry of such an element is rather too specialized for inclusion in a survey book of this type. The plan followed in this chapter will be to discuss some of the topics of sulfur chemistry separately from those of selenium and tellurium because in several regards sulfur is somewhat different from the other two elements.

As in the case of most chemical materials known in ancient times, sulfur or *brimstone* is sometimes found in its free state. It was one of the earliest chemicals known to man, and it is probable that it was recognized at least as long ago as 2000 BC. Its medicinal use was known even to the ancient Romans and Egyptians. The name sulfur is derived from the Sanskrit, *sulveri*, and the Latin, *sulphurium*.

## 16.1 OCCURRENCE OF SULFUR

Sulfur-containing compounds and materials are quite common in the volcanic regions of the earth. These include Italy, Sicily, Greece, and the Middle East (Palestine, Iran, and others) as well as Mexico, Chile, Peru, Japan, and the United States. All of these areas of the world have or have had some volcanic activity. There are also several important ores that are sources of metals in which the metals are present as the sulfides. Some of these are *galena* (PbS), *zinc blende* (ZnS), *cinnabar* (HgS), *iron pyrites* (FeS<sub>2</sub>), *gypsum* (CaSO<sub>4</sub>), and *chalcopyrite* (CuS<sub>2</sub>).

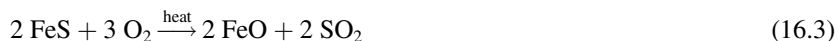
The fact that sulfur is often found in areas of volcanic activity may be due to the reaction of sulfur dioxide, SO<sub>2</sub>, and hydrogen sulfide, H<sub>2</sub>S, both gases being found in volcanic emissions. One such reaction can be represented by the equation



Although elemental sulfur exists primarily as S<sub>8</sub> rings, for convenience it will be represented simply as S in most equations. Sulfur dioxide may be produced by burning sulfur,



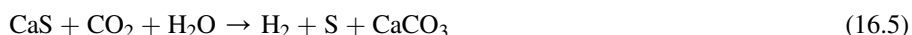
or the roasting of metal sulfides,



Hydrogen sulfide is produced by the action of acids on sulfide minerals. For example, the reaction of galena with an acid to produce hydrogen sulfide can be represented as

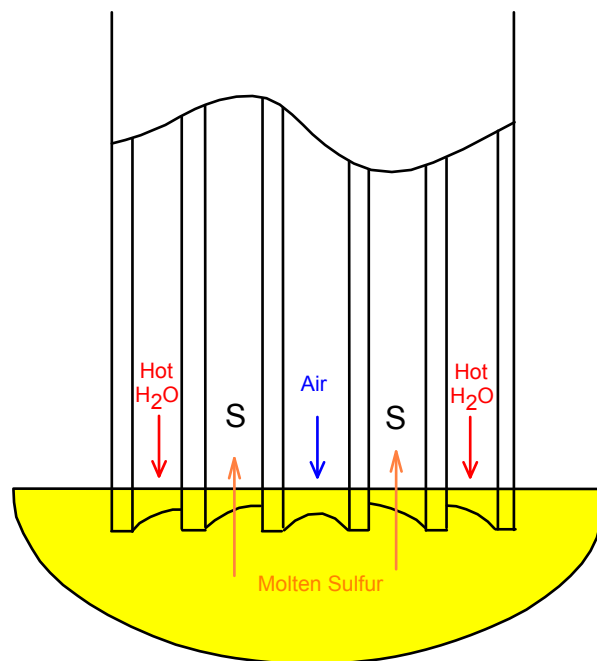


The reaction of calcium sulfide with H<sub>2</sub>O and CO<sub>2</sub> also produces sulfur.



The CaS is produced in nature by the reduction of CaSO<sub>4</sub>, possibly by bacterial action, and this process may be responsible for the production of sulfur from native CaSO<sub>4</sub> (*gypsum*). The processes described above show that the presence of sulfur can be accounted for by processes that are naturally occurring.

Areas of the earth that are rich in sulfur minerals also contain *limestone*, CaCO<sub>3</sub>, and *gypsum*, CaSO<sub>4</sub>. These impure ores may contain up to as much as 25% elemental sulfur. The sulfur can be separated by heating the mixture, and, because sulfur melts at about 120 °C, the liquid sulfur flows away from the solid rocky materials to a lower portion of the sloping floor in a furnace. Because sulfur readily burns with the liberation of a great amount of heat, it can even be used as the fuel for the process.



**FIGURE 16.1** The Frasch process for obtaining sulfur.

In the United States, sulfur is found mainly in Texas and Louisiana where it is obtained by the *Frasch process* that can be illustrated as shown in [Figure 16.1](#). This process makes use of three concentric pipes. The deposits are found several hundred feet below the surface and are heated to melting by sending water that has been heated under pressure to 170 °C down the outside pipe. Compressed air is forced down the center pipe, and the molten sulfur is then forced by pressure up the third pipe along with water and air.

In present times, the most important use of sulfur is in making sulfuric acid, and although  $\text{H}_2\text{SO}_4$  is now the most important chemical containing sulfur, this has not always been so. In fact, it is believed that  $\text{H}_2\text{SO}_4$  was discovered only in about the tenth century. To the ancients, sulfur and its combustion product,  $\text{SO}_2$ , were more important. Elemental sulfur is one of the components of black gunpowder, so its importance in this area was enormous until the late 1800s when smokeless powder (nitrocellulose) was developed.

The discussion up to this point has been aimed at putting sulfur chemistry in its historical and economic perspective. In some ways, the chemistry of sulfur and its compounds has been of paramount importance for centuries, a trend that continues to the present time. The major use of sulfur today is in the production of sulfuric acid, and approximately two-thirds of the sulfur produced is used in this way. Over half of the sulfuric acid produced is used in making fertilizers. Other uses of sulfur compounds will be described in the following sections.

## 16.2 OCCURRENCE OF SELENIUM AND TELLURIUM

Tellurium was discovered in 1782 by F. J. Müller von Reichenstein, and selenium was discovered in 1817 by J. J. Berzelius. Polonium was discovered in 1898 by Marie Curie, but of the nearly 30 known isotopes of polonium, none are stable so the element is of much less importance chemically.

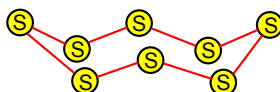
Tellurium is named from the Latin, *tellus*, meaning “earth.” Selenium is named from the Greek, *selene*, meaning “moon.” Apparently, these names were chosen to emphasize the fact that the elements are usually found together. Polonium was named in honor of Poland, Marie Curie’s native country. Selenium and tellurium are trace elements and generally occur with the corresponding sulfur compounds. Some selenium is contained in native sulfur found in Sicily. Selenium and tellurium are also obtained as by-products from the processing of sulfur-containing materials. The largest amount of Se and Te produced in the United States is obtained from the sludge produced in the electrolytic refining of copper. That anode sludge or slime is also the source of some silver and gold.

### 16.3 ELEMENTAL SULFUR

Sulfur melts at approximately 120 °C so it can be melted by high pressure steam which allows sulfur to be obtained by the *Frasch process* described earlier. Solid sulfur exists in several allotropic forms. At room temperature, the rhombic crystalline form is stable whereas above 105 °C, a monoclinic form is stable. These two crystalline forms are shown in Figure 16.2.

Other forms of sulfur include a plastic or amorphous form that can be obtained by rapidly cooling molten sulfur. For example, pouring molten sulfur at 160 °C into water produces the amorphous form.

In the solid state, the species that is most stable is the S<sub>8</sub> molecule that has the structure of a puckered ring.



In the gas phase, the paramagnetic S<sub>2</sub> molecules exist, but a rather large number of other molecular species are also known. Although sulfur is insoluble in H<sub>2</sub>O, it is soluble in nonpolar organic liquids such as CS<sub>2</sub> and C<sub>6</sub>H<sub>6</sub>. The molecular weight determined from cryoscopic measurements on solutions of sulfur agrees with the S<sub>8</sub> formula. However, sulfur also exists as rings that contain 6, 7, 9, 10, 11, 12, 18, and 20 atoms.

The properties of liquid sulfur are unusual in many ways, and the variation in viscosity with temperature is shown in Figure 16.3. An explanation for this behavior lies in the nature of the actual species that are present in the liquid. As the temperature is raised above the melting point, some S<sub>8</sub> molecules are broken and form larger aggregates, and probably both linear and ring structures are present. The larger molecules do not flow as readily, and because more S<sub>8</sub> units are disrupted at higher temperatures, the viscosity increases with increasing temperature. After a certain temperature (about 170–190 °C), these larger aggregates dissociate to produce smaller fragments and the viscosity decreases as the temperature is increased.

In the vapor state, sulfur contains several species that are in equilibrium. Among them are S<sub>8</sub>, S<sub>6</sub>, S<sub>4</sub>, and S<sub>2</sub> with the relative amounts depending on the temperature and pressure. Some of the properties of elemental sulfur, selenium, and tellurium are shown in Table 16.1.

Elemental sulfur reacts with many elements, both metallic and nonmetallic. Heavy metals readily form sulfides in preference to oxides in accord with the hard–soft interaction principle (see Chapter 6). Some examples of this behavior are illustrated by the following equations.



Typical reactions with nonmetals are as follows.

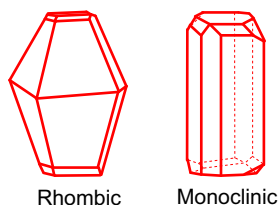


FIGURE 16.2 Crystalline forms of sulfur.

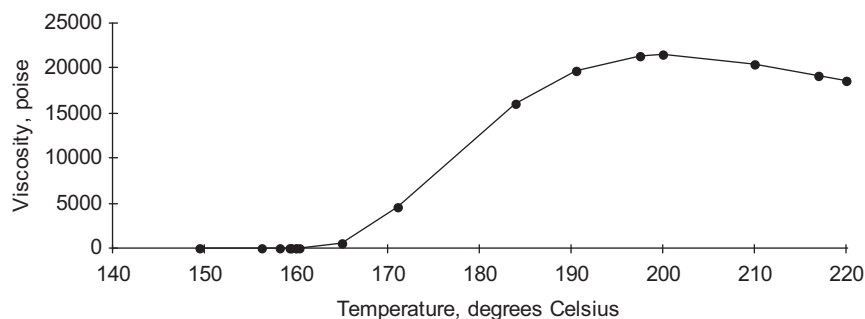


FIGURE 16.3 The viscosity of sulfur as a function of temperature.

TABLE 16.1 Properties of Sulfur, Selenium, and Tellurium

Property	S	Se	Te
m.p., °C	119	220	450
b.p., °C	441	688	1390
$\Delta H_{\text{vap}}$ , kJ mol <sup>-1</sup>	9.6	26.4	50.6
Density, gm cm <sup>-3</sup>	2.06	4.82	6.25
First ionization potential, eV	10.38	9.75	9.01
Electronegativity	2.6	2.4	2.1
X–X bond energy, kJ mol <sup>-1</sup>	213	330	260
Atomic radius, pm	104	117	137

Another interesting property of elemental sulfur is its ability to add to atoms that are already bonded. For example, thiosulfates are prepared by the reaction of sulfites with sulfur in boiling solutions.



Phosphorus compounds also show the ability to add sulfur as illustrated by the equation



Sulfur can be oxidized to  $\text{SO}_2$  by hot concentrated  $\text{HNO}_3$ .



Thiocyanates are obtained by the reaction of sulfur with cyanides.



Reactions of this type appear to be consistent with the opening of the  $\text{S}_8$  ring as being the rate determining step. Another interesting reaction of sulfur is that of removing hydrogen from organic compounds. Heating paraffin with sulfur readily produces  $\text{H}_2\text{S}$  by removal of hydrogen atoms, which results in the formation of carbon–carbon double bonds. Finally, it should be noted that sulfur dissolves in solutions containing  $\text{S}^{2-}$  to form polysulfides.



## 16.4 ELEMENTAL SELENIUM AND TELLURIUM

A major use of selenium has been in photoelectric devices. Its conductivity increases with illumination, and this provides a way of measuring light intensity or operating electrical switches. However, newer types of photocells are available that are made of other materials (such as cadmium sulfide). A second and more important use of selenium is in rectifiers to convert

alternating current to direct current. Also, some pigments contain selenium and tellurium compounds, and both elements have been used in vulcanization of rubber. Selenium compounds have been used in dandruff treatment shampoos, and low levels of selenium may be necessary for dietary balance. Some studies have shown that persons whose diets are deficient in selenium may have a higher incidence of heart attacks.

Although sulfur shows a great tendency toward catenation, selenium and tellurium show much less tendency in this regard. Some structures containing chains and rings of Se and Te atoms are known. At high temperatures, selenium vapor contains  $\text{Se}_8$ ,  $\text{Se}_6$ , and  $\text{Se}_2$  with the amount of each depending on the temperature. In tellurium vapor, the predominant species is  $\text{Te}_2$ . Unlike the behavior described earlier for liquid sulfur, the viscosity of liquid selenium decreases with increasing temperature. Several modifications of solid selenium exist with the most stable being a gray or metallic form. A black vitreous form results when liquid selenium is cooled, and the element is available commercially in this form. Heating vitreous selenium to approximately  $180^\circ\text{C}$  results in it being transformed into the gray form.

Tellurium (m.p.  $450^\circ\text{C}$ , density  $6.25\text{ g cm}^{-3}$ ) is more metallic in its appearance, but it is not a good electrical conductor as are most metals. Polonium, on the other hand, is typically metallic in its electrical properties. Selenium and tellurium are best regarded as semiconductors, and sulfur is nonmetallic in behavior (an insulator). Thus, the usual trend from nonmetallic to metallic behavior is shown in going down Group VIA of the periodic table. All of these elements differ substantially from oxygen in their chemical properties.

## 16.5 REACTIONS OF ELEMENTAL SELENIUM AND TELLURIUM

Selenium and tellurium are rather reactive elements, and they combine readily with a variety of metals and nonmetals. For example, they readily burn to produce the dioxides.



Reactions with metals produce selenides and tellurides as illustrated by the following reactions.



Reactions with halogens can produce a variety of products depending on the relative amounts of the reactants and the reaction conditions.



Although the reaction of selenium with hydrogen does not take place readily, the elements do react at high temperature. Tellurium does not react with hydrogen, and both  $\text{H}_2\text{Se}$  and  $\text{H}_2\text{Te}$  are less stable than  $\text{H}_2\text{S}$ . Hot concentrated  $\text{HNO}_3$  will dissolve both selenium and tellurium as shown in the following equation.



The elements also react with cyanide ion to form the selenocyanate and tellurocyanate ions.



Cationic species containing selenium and tellurium can be obtained by the following reactions.



The ions  $\text{S}_4^{2+}$ ,  $\text{Se}_4^{2+}$ , and  $\text{Te}_4^{2+}$  all appear to have square planar structures in which the bond length is short enough to indicate multiple bonding. The ions  $\text{S}_8^{2+}$ ,  $\text{Se}_8^{2+}$ ,  $\text{Te}_8^{2+}$  all have puckered ring structures.

## 16.6 HYDROGEN COMPOUNDS

The elements S, Se, and Te all form hydrogen compounds. By far the most common are H<sub>2</sub>S, H<sub>2</sub>Se, and H<sub>2</sub>Te, although some unstable compounds such as H<sub>2</sub>S<sub>2</sub> (m.p.  $-88\text{ }^{\circ}\text{C}$ , b.p.  $74.5\text{ }^{\circ}\text{C}$ ) and H<sub>2</sub>S<sub>6</sub> are known. These compounds contain sulfur atoms bonded together in chains. Some of the properties of the H<sub>2</sub>X compounds are shown in Table 16.2.

The boiling point of H<sub>2</sub>O is abnormally high due to intermolecular hydrogen bonding, but there is little if any hydrogen bonding in H<sub>2</sub>S, H<sub>2</sub>Se, or H<sub>2</sub>Te. The heats of formation of the compounds (shown in Table 16.2) indicate the decrease in stability arising from the mismatch in orbital size between these atoms and hydrogen.

All of these H<sub>2</sub>X compounds are about as toxic as HCN, and many deaths have been caused by H<sub>2</sub>S poisoning. The odor is noticeable at concentrations as low as 0.025 ppm and gets progressively worse as the concentration increases. However, at high concentrations (200–700 ppm) the objectionable odor largely disappears because the sense of smell is affected. At these concentrations, the odor is actually somewhat sweet, and it is possible that some accidental deaths have occurred because the odor became less unpleasant and the danger was not recognized.

Hydrogen selenide can be prepared by direct reaction of the elements, but owing to the instability of H<sub>2</sub>Te, it is not prepared by direct combination. Both H<sub>2</sub>Se and H<sub>2</sub>Te can be prepared by acid hydrolysis of selenides and tellurides, respectively. The process can be shown as follows.



Hydrolysis of sulfides produces H<sub>2</sub>S, but this gas can also be prepared by reaction of the elements. All of these hydrogen compounds are acidic, although only the first step of the dissociation in water is significant.



From the dissociation constants shown in Table 16.2, it is apparent that the strengths of the compounds as acids varies as H<sub>2</sub>S < H<sub>2</sub>Se < H<sub>2</sub>Te. This can be explained in terms of the less effective overlap of the 1s orbital of hydrogen with the *p* orbital on the Group VIA element. An interesting series of bond angles is shown by these C<sub>2v</sub> molecules. In H<sub>2</sub>O, the lone pair repulsion causes the bonding pairs to be forced inward (to about 104.5°) from the usual tetrahedral bond angle. For S, Se, and Te, it appears that effectiveness of overlap is not enhanced by hybridization so that the overlap involves essentially pure *p* orbitals on the S, Se, and Te atoms. In fact, it appears that using essentially pure *p* orbitals and allowing for slight lone pair repulsion should produce a bond angle of almost exactly 90°, as is the case for H<sub>2</sub>Se and H<sub>2</sub>Te.

Although the usual H<sub>2</sub>X compounds of sulfur, selenium, and tellurium are the most important ones, the sulphanes deserve some consideration. These molecules have the general formula H<sub>2</sub>S<sub>*x*</sub>, and they owe their existence to the ability of sulfur to bond to itself. Naming the compounds is accomplished by indicating the number of sulfur atoms; H<sub>2</sub>S<sub>3</sub> is trisulfane, H<sub>2</sub>S<sub>6</sub> is hexasulfane, etc. Some of these compounds can be prepared by acidifying a solution containing

**TABLE 16.2** Properties of Hydrogen Compounds of Group VIA Elements

Property	H <sub>2</sub> O	H <sub>2</sub> S	H <sub>2</sub> Se	H <sub>2</sub> Te
m.p., °C	0.00	-85.5	-65.7	-51
b.p., °C	100.0	-60.7	-41.3	-2.3
$\Delta H_f^\circ$ (gas), kJ mol <sup>-1</sup>	-285.9	+20.1	+73.1	+99.6
Dipole moment, Debye	1.85	0.97	0.62	0.2
Acid $K_{a1}$	$1.07 \times 10^{-16}$	$1.0 \times 10^{-9}$	$1.7 \times 10^{-4}$	$2.3 \times 10^{-3}$
Acid $K_{a2}$	—	$1.2 \times 10^{-15}$	$1.0 \times 10^{-10}$	$1.6 \times 10^{-11}$
Bond angle, degrees	104.5	92.3	91.0	89.5
H–X bond energy, kJ mol <sup>-1</sup>	464	347	305	268



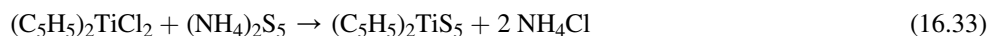
polysulfide species. Compounds having the formula  $\text{H}_2\text{S}_n$  ( $n > 2$ ) can be obtained in this way. Higher sulfanes result from the reaction of  $\text{H}_2\text{S}_2$  and  $\text{S}_2\text{Cl}_2$ .



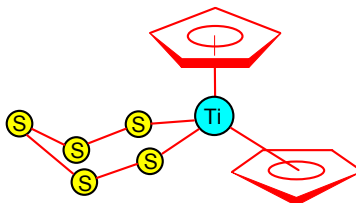
Polysulfide salts of several metals are known, and they can be prepared by the reaction of metals with sulfur in liquid ammonia or by dissolving sulfur in the molten metal sulfide. The sulfanes are all rather unstable and decompose easily to produce  $\text{H}_2\text{S}$ .



Some of the polysulfides undergo interesting and unusual reactions such as that shown for  $\text{S}_5^{2-}$ .



The polysulfide group can be regarded as a ligand with a  $-2$  charge giving a product having the structure shown below.



In this structure, the cyclopentadiene rings are bonded to Ti by donation of the  $\pi$  electrons. Organic compounds containing  $-\text{S}-\text{S}-$  linkages are of considerable importance, and compounds of this type also occur in biological materials. However, there is not a corresponding series of Se and Te compounds, and even  $\text{H}_2\text{Se}$  and  $\text{H}_2\text{Te}$  are not especially stable.

## 16.7 OXIDES OF SULFUR, SELENIUM, AND TELLURIUM

Although the molecules  $\text{S}_2\text{O}$  and  $\text{SO}$  are known, they are of no real importance. The existence of similar species containing Se and Te is questionable. The only oxides of importance are the dioxides and trioxides, and consequently, only these compounds will be discussed here.

### 16.7.1 Dioxides

All of the elements in Group VIA form well-defined dioxides, and Table 16.3 shows some of the properties of these compounds.

Gaseous  $\text{SO}_2$  is a toxic gas with a sharp odor. It is produced in a large number of reactions, but the usual method is burning sulfur in air.



Selenium and tellurium react in a similar way. Sulfur dioxide is also released when sulfites react with strong acids.



Selenium dioxide and tellurium dioxide are also produced by treating the elements with hot concentrated nitric acid and driving off water to leave a solid residue. Metallic tin is converted to  $\text{SnO}_2$  in a similar process (see Chapter 12).

**TABLE 16.3** Properties of Group VIA Dioxides

Property	$\text{SO}_2$	$\text{SeO}_2$	$\text{TeO}_2$
m.p., $^\circ\text{C}$	-75.5	340	733
b.p., $^\circ\text{C}$	-10.0	Subl.	—
$\Delta H_f^\circ$ , $\text{kJ mol}^{-1}$	-296.9	-230.0	-325.3
$\Delta G_f^\circ$ , $\text{kJ mol}^{-1}$	-300.4	-171.5	-269.9

Large amounts of SO<sub>2</sub> are produced when metal sulfide ores are roasted in air. *Pyrite*, FeS<sub>2</sub>, is an important ore of iron in many parts of the world. The first step in the processing of this ore is roasting the sulfide to produce the oxide,



Large quantities of SO<sub>2</sub> are also produced by the burning of high-sulfur coal as a fuel. If released into the atmosphere, this SO<sub>2</sub> reacts with water to form acid rain. A great deal of study is directed toward removing and utilizing the SO<sub>2</sub> from stack gases and the conversion of the SO<sub>2</sub> into sulfuric acid. This source of sulfuric acid has decreased the amount of acid produced by the burning of elemental sulfur and has helped in the prevention of air pollution.

Historically, SO<sub>2</sub>, because of its reducing character, has been used as an atmosphere to retard such oxidation processes as the spoilage of fruit. It is probable that sulfur was burned by the ancients to produce this “preservative” and purifying agent.

Liquid SO<sub>2</sub> has a number of important uses. It has been used as a refrigerant, although ammonia and chlorofluorocarbons have also been used for this purpose. Liquid SO<sub>2</sub> has been extensively utilized as a nonaqueous solvent (see Chapter 6). Its low electrical conductivity is more likely to be due to the presence of trace amounts of impurities than to a very slight degree of ionization. If ionization of liquid SO<sub>2</sub> did occur, it could be represented as



in which SO<sup>2+</sup> is the acidic species and SO<sub>3</sub><sup>2-</sup> is the basic species. However, SO<sub>2</sub> is much more likely to react in most cases by complex formation than by ionization. This is especially true because SO<sub>2</sub> can act as either an electron pair acceptor,



or an electron pair donor,



It also forms some complexes with metals, particularly the second and third row transition metals in low oxidation states.

Sulfur dioxide reacts with chlorine to yield sulfuryl chloride, O<sub>2</sub>SCl<sub>2</sub>, (commonly written as SO<sub>2</sub>Cl<sub>2</sub>).

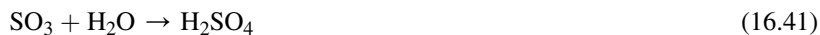


This reaction is carried out on a large scale in which activated charcoal or FeCl<sub>3</sub> is used as a catalyst. The S–Cl bonds in O<sub>2</sub>SCl<sub>2</sub> are very reactive so it is a useful reagent for synthesizing numerous other compounds.

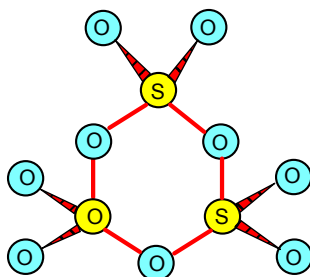
Because of its π electron system and its polarizability, liquid SO<sub>2</sub> is a good solvent for aromatic hydrocarbons. Aliphatic compounds are less soluble in liquid SO<sub>2</sub>, so it is possible to devise a solvent extraction process utilizing liquid SO<sub>2</sub> to separate aliphatic and aromatic hydrocarbons.

### 16.7.2 Trioxides

Sulfur trioxide is produced by the oxidation of SO<sub>2</sub>. Although the oxidation of SO<sub>2</sub> is catalyzed by platinum, arsenic or halogens strongly poison the catalyst. Another suitable catalyst is sodium vanadate made from V<sub>2</sub>O<sub>5</sub> and Na<sub>2</sub>O. At an operating temperature of 400 °C, the catalyst is a liquid. The importance of SO<sub>3</sub> is largely related to the fact that it is the anhydride of sulfuric acid,

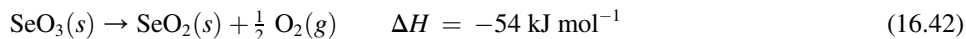


Three solid phases of SO<sub>3</sub> are known. One form, often referred to as γ-SO<sub>3</sub>, consists of cyclic trimers that have the following structure.

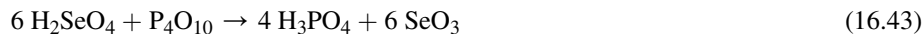


A second form,  $\beta$ -SO<sub>3</sub>, consists of polymeric molecules in helical chains made up of SO<sub>4</sub> tetrahedra. The third form,  $\alpha$ -SO<sub>3</sub>, has the lowest vapor pressure and the highest melting point of the three forms. In this form, chains are present, but cross-linking produces a layered structure.

Selenium trioxide is less stable than the dioxide so it dissociates readily.

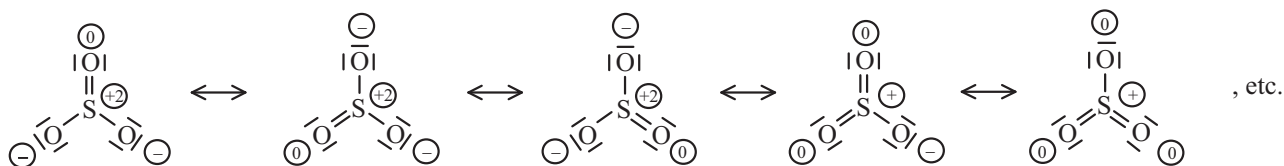


The trioxide can be prepared by dehydration of H<sub>2</sub>SeO<sub>4</sub>.



A variety of organic solvents will dissolve SeO<sub>3</sub>. Tellurium trioxide can be obtained by dehydrating Te(OH)<sub>6</sub>, and the solid compound exists in two forms. Neither SeO<sub>3</sub> nor TeO<sub>3</sub> is of significant commercial importance although both are strong oxidizing agents.

In the gas phase, the SO<sub>3</sub> molecule has a trigonal planar structure ( $D_{3h}$ , dipole moment = 0). The S–O distance in SO<sub>3</sub> is 141 pm and is shorter than a typical single bond length indicating a significant amount of multiple bonding. A large number of resonance structures showing this multiple bonding can be drawn.



The first three of these structures have a +2 formal charge on the sulfur atom so other structures are included that reduce this charge. The contributions by all of the structures indicate that there is a positive formal charge on the sulfur atom which is consistent with the strong Lewis acidity of that site.

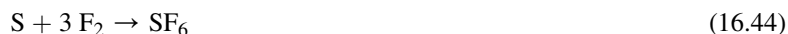
Several important halogen derivatives of sulfuric acid are produced from SO<sub>3</sub>. Some of these will be described in the next section, and some are discussed with the chemistry of sulfuric acid.

## 16.8 HALOGEN COMPOUNDS

The halogen compounds of sulfur include some compounds that have considerable utility in industrial applications. With a single exception, SF<sub>6</sub>, they are reactive compounds that undergo reactions characteristic of nonmetal halides. The tetrachlorides and tetrafluorides are the most important of the halogen compounds of the Group VIA elements. Table 16.4 shows the melting points and boiling points of the well-characterized compounds containing Group VIA elements and halogens.

Sulfur hexafluoride sublimates at  $-64^\circ\text{C}$  to produce a dense gas ( $6.14 \text{ g L}^{-1}$ ). Under a pressure of 2 atm, the melting point is  $-51^\circ\text{C}$ . The molecule has the expected octahedral structure and a dipole moment of zero. The compound is so inert that it is used as a gaseous insulator, and rats allowed to breathe a mixture of SF<sub>6</sub> and oxygen show no ill effects after several hours of exposure. This inertness is a result of the molecule having no vacant bonding site or unshared electron pairs on sulfur to initiate a reaction and the fact that six fluorine atoms shield the sulfur atom from attack. Consequently, there is no low energy pathway for reactions to occur, and the compound is inert even though many reactions are thermodynamically favored.

Sulfur hexafluoride can be prepared by direct combination of the elements, although small amounts of SF<sub>4</sub> and S<sub>2</sub>F<sub>10</sub> also result.



The fluorination can also be carried out using ClF<sub>3</sub> or BrF<sub>5</sub> as the fluorinating agent.



Disulfur decafluoride, S<sub>2</sub>F<sub>10</sub>, can be prepared by the photochemical reaction of SF<sub>5</sub>Cl with H<sub>2</sub>.



**TABLE 16.4** Halogen Compounds of the Group VIA Elements

Compound	m.p., °C	b.p., °C
S <sub>2</sub> F <sub>2</sub>	-133	15
SF <sub>4</sub>	-121	-38
SF <sub>6</sub>	-51 (2 atm)	-63.8 (subl)
S <sub>2</sub> F <sub>10</sub>	-52.7	30
SCl <sub>2</sub>	-122	59.6
S <sub>2</sub> Cl <sub>2</sub>	-82	137.1
S <sub>3</sub> Cl <sub>2</sub>	-45	—
SCl <sub>4</sub>	-31 (dec)	—
S <sub>2</sub> Br <sub>2</sub>	-46	90 (dec)
SeF <sub>4</sub>	-9.5	106
SeF <sub>6</sub>	-34.6 (1500 torr)	-34.8 (945 torr)
Se <sub>2</sub> Cl <sub>2</sub>	-85	127 (733 torr)
SeCl <sub>4</sub>	191 (subl)	—
Se <sub>2</sub> Br <sub>2</sub>	-146	225 d
TeF <sub>4</sub>	129.6	194 d
TeF <sub>6</sub>	-37.8	-38.9 (subl)
Te <sub>2</sub> F <sub>10</sub>	-33.7	59
TeCl <sub>2</sub>	208	328
TeCl <sub>4</sub>	224	390
TeBr <sub>4</sub>	380	414 (dec)

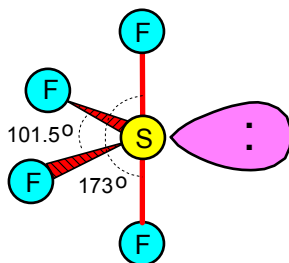
As in the case of SF<sub>6</sub>, the compound is unreactive, but its reactivity is increased at higher temperatures, probably as the result of some dissociation.



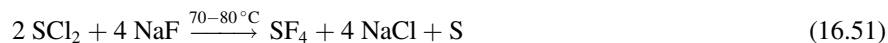
In general, SeF<sub>6</sub> and TeF<sub>6</sub> are similar to SF<sub>6</sub>, but they are more reactive. The increased reactivity reflects the fact that the larger Se and Te atoms are less crowded by six fluorine atoms than are smaller S atoms. Both SeF<sub>6</sub> and TeF<sub>6</sub> slowly hydrolyze in water.



Sulfur tetrafluoride is a reactive gas (b.p. -40 °C), and the polar molecule has a C<sub>2v</sub> structure (dipole moment = 0.63 D) that can be shown as follows.



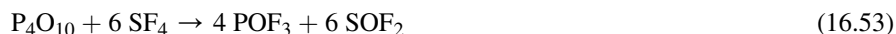
It can be prepared by several reactions, two of which are as follows.



Unlike SF<sub>6</sub>, SF<sub>4</sub> is toxic and readily hydrolyzes.



It also functions as a reactive fluorinating agent toward many compounds, and typical reactions include the following.



Because the SF<sub>4</sub> molecule has an unshared pair of electrons on the sulfur atom, it can behave as a Lewis base. The reaction with the strong Lewis acid SbF<sub>5</sub> can be shown as



As a result of sulfur not being restricted to a maximum of 8 or 10 electrons in the valence shell, SF<sub>4</sub> can also act as an electron pair acceptor, so it can behave as a Lewis acid also.

The SeF<sub>4</sub> molecule has a distorted structure similar to that of SF<sub>4</sub>, and its dipole moment is 1.78 D. Chemically, the tetrafluorides of Se and Te behave in much the same way as SF<sub>4</sub>. Selenium tetrafluoride can be prepared by direct combination of the elements or by the reaction of SF<sub>4</sub> with SeO<sub>2</sub> at elevated temperatures. Tellurium tetrafluoride can be prepared by the reaction of SeF<sub>4</sub> and TeO<sub>2</sub> at 80 °C.

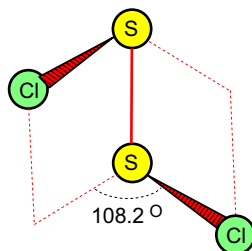
A number of chlorides are known for the elements S, Se, and Te. For example, elemental sulfur reacts with chlorine to give S<sub>2</sub>Cl<sub>2</sub>,



although SCl<sub>2</sub> can also be produced by this reaction if excess Cl<sub>2</sub> is used.



The structure of S<sub>2</sub>Cl<sub>2</sub> can be shown as



Because of the ability of sulfur to bond to itself in chains, other ill-defined products result from the reaction of sulfur with S<sub>2</sub>Cl<sub>2</sub>.



Compounds having the formula S<sub>n</sub>Cl<sub>2</sub> can be considered as derivatives of the sulfanes, and the higher chlorosulfanes can be prepared by the reaction of S<sub>2</sub>Cl<sub>2</sub> with sulfanes.



Two other important compounds, SOCl<sub>2</sub> and SO<sub>2</sub>Cl<sub>2</sub>, can be obtained by oxidation of SCl<sub>2</sub>.

Unlike  $\text{SCl}_4$  which is unstable, both  $\text{SeCl}_4$  and  $\text{TeCl}_4$  are stable. There is some dissociation of  $\text{SeCl}_4$  in the vapor state, but  $\text{TeCl}_4$  is stable up to  $500^\circ\text{C}$ . Both molecules have a structure similar to that of  $\text{SF}_4$ , and  $\text{TeCl}_4$  has a dipole moment of 2.54 D. Molten  $\text{TeCl}_4$  is a good electrical conductor, possibly due to the ionization reaction



## 16.9 NITROGEN COMPOUNDS

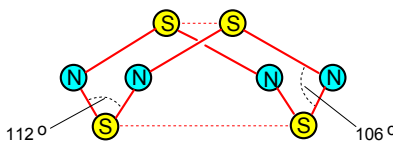
A considerable number of sulfur–nitrogen compounds exist, and some of them have unusual properties. The most common compound of this type is  $\text{S}_4\text{N}_4$ , tetrasulfur tetranitride. The compound can be prepared by the following reaction.



It may also be prepared by the reaction of gaseous  $\text{NH}_3$  with an ether solution of  $\text{S}_2\text{Cl}_2$ .

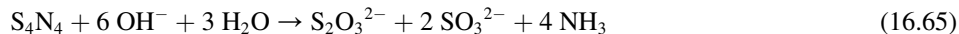


Tetrasulfur tetranitride is thermochromic. The compound is nearly colorless at  $-190^\circ\text{C}$ , orange at  $25^\circ\text{C}$ , and dark red at  $100^\circ\text{C}$ . The compound can be purified by sublimation at  $100^\circ\text{C}$  under reduced pressure. Although the compound is stable under a variety of conditions, explosions result when the compound is subjected to shock, which is not surprising considering that it has a heat of formation of  $+538.9 \text{ kJ mol}^{-1}$ . A number of organic solvents will dissolve  $\text{S}_4\text{N}_4$  including such common solvents as benzene, carbon disulfide, chloroform, and carbon tetrachloride. The structure of  $\text{S}_4\text{N}_4$  consists of a ring in which the four nitrogen atoms are coplanar.



There is apparently significant interaction between the sulfur atoms. The experimental S–S distance is about 258 pm (that of an S–S single bond is about 210 pm), and the S–N distance is about 162 pm. The sum of the van der Waals radii for two sulfur atoms is about 370 pm, but the S–S distance where an N atom is linking them is only about 271 pm. Thus, an extended  $\pi$  electron system appears to be involved in the bonding. A sizable number of resonance structures can be drawn in which there are unshared pairs of electrons on the N and S atoms, but a molecular orbital approach that takes into account some delocalization of electrons over the entire structure provides a more satisfactory approach to the bonding.

Tetrasulfur tetranitride undergoes a number of interesting reactions. Hydrolysis of the compound in basic solution produces  $\text{NH}_3$  as the nitrogen-containing product in accord with nitrogen having the higher electronegativity.



It also reacts with dry ammonia to produce an ammoniate.



Halogen derivatives are obtained by reactions with  $\text{Cl}_2$  and  $\text{Br}_2$ , and the reaction with chlorine can be represented as follows.



The product,  $\text{S}_4\text{N}_4\text{Cl}_4$ , maintains the ring structure of  $\text{S}_4\text{N}_4$  but has the halogen atoms attached to the sulfur atoms.

With Lewis acids,  $\text{S}_4\text{N}_4$  undergoes reactions in which it functions as an electron pair donor to form adducts without ring opening.



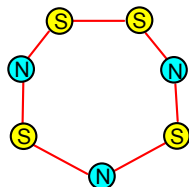
Reduction of  $\text{S}_4\text{N}_4$  occurs when it reacts with  $\text{SnCl}_2$  in ethanol.



Similarly, reaction with  $\text{AgF}_2$  in  $\text{CCl}_4$  fluorinates the ring to yield  $(\text{F-SN})_4$  in which the fluorine atoms are attached to the sulfur atoms. Other ring structures can be obtained by opening or reorienting the  $\text{S}_4\text{N}_4$  ring. For example, the reaction with  $\text{SOCl}_2$  is interesting in this connection because it produces the thiotrithiazyl,  $\text{S}_4\text{N}_3^+$ , ring.



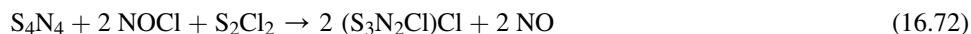
The thiotrithiazyl name is derived from the fact that the S–N unit is the thiazyl group and  $\text{S}_4\text{N}_3^+$  contains three of these units plus one sulfur (thio) atom. The  $\text{S}_4\text{N}_3^+$  ring is planar and has the arrangement of atoms shown below.



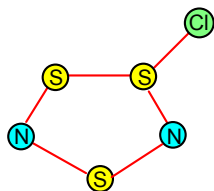
This ring also results from the reaction



Thiotrithiazyl chloride explodes when heated. Tetrasulfur tetranitride reacts with  $\text{NOCl}$  in the following way to yield a product that contains the  $\text{S}_3\text{N}_2\text{Cl}^+$  cation.



The  $\text{S}_3\text{N}_2\text{Cl}^+$  cation consists of a five-membered ring having Cl attached to a sulfur atom.



Disulfur dinitride ( $\text{S}_2\text{N}_2$ ) can be obtained by passing  $\text{S}_4\text{N}_4$  vapor through silver wool at  $300^\circ\text{C}$ ; however,  $\text{S}_2\text{N}_2$  explodes easily. The structure of the monomer contains a four-membered ring with alternating sulfur and nitrogen atoms. It is a Lewis base and forms adducts with a variety of Lewis acids. A great deal of interest has recently been focused on this compound because of its ability to polymerize to give high molecular weight polymers,  $(\text{SN})_x$ , that have metallic properties. They are malleable, and the material becomes a superconductor at  $0.3 \text{ K}$ . The polymerization can be carried out by keeping the compound at room temperature for 3 days and then completing the process by heating the material to  $75^\circ\text{C}$  for 2 h.

A number of other sulfur–nitrogen compounds are known. Generally, they contain at least one other type of atom and are prepared from  $\text{S}_4\text{N}_4$ . For example, tetrasulfur tetraimide is obtained by the reduction of  $\text{S}_4\text{N}_4$ . In fact, a series of sulfur imides having the formula  $\text{S}_n(\text{NH})_{8-n}$  are known, as are some halogen derivatives.

The selenium and tellurium compounds of nitrogen are much less numerous than the sulfur compounds. Selenium forms the compound  $\text{Se}_4\text{N}_4$ , but tellurium gives a compound that has a composition between  $\text{Te}_3\text{N}_4$  and  $\text{Te}_4\text{N}_4$ . Both the selenium and tellurium compounds are dangerously explosive. A series of polymeric selenium–nitrogen materials containing halogens, e.g.,  $(\text{Se}_2\text{N}_2\text{Br})_n$ , is known.

## 16.10 OXYHALIDES OF SULFUR AND SELENIUM

Sulfur and selenium form similar series of compounds containing oxygen and halogens. These compounds are reactive and usually undergo reactions in which sulfur–halogen or selenium–halogen bonds are broken. Because of this, they are useful for preparing many compounds containing sulfur–carbon and selenium–carbon bonds. Properties of some of these oxyhalide compounds are shown in Table 16.5.

**TABLE 16.5** Properties of Oxyhalides of Sulfur and Selenium

Compound	m.p., °C	b.p., °C	$\mu$ , D
SOF <sub>2</sub>	-110.5	-43.8	—
SOCl <sub>2</sub>	-106	78.8	1.45
SOBr <sub>2</sub>	-52	183	—
SOCIF	-139.5	12.1	—
SeOF <sub>2</sub>	4.6	124	—
SeOCl <sub>2</sub>	8.6	176.4	—
SeOBr <sub>2</sub>	41.6	dec.	—
SO <sub>2</sub> F <sub>2</sub>	-136.7	-55.4	1.12
SO <sub>2</sub> Cl <sub>2</sub>	-54.1	69.1	1.81
SO <sub>2</sub> ClF	-124.7	7.1	1.81

### 16.10.1 Oxidation State +4

In this series of sulfur compounds, the most important members are SOCl<sub>2</sub> and SOF<sub>2</sub>. Thionyl chloride, SOCl<sub>2</sub>, which is shown more correctly as OSCl<sub>2</sub> because the oxygen atom is bonded to the sulfur atom, can be prepared by the reaction of SO<sub>2</sub> with a strong chlorinating agent such as PCl<sub>5</sub>.



It can also be prepared by the reaction of lower sulfur halides such as S<sub>2</sub>Cl<sub>2</sub> or SCl<sub>2</sub> with SO<sub>3</sub> as illustrated by the following equation.



Thionyl bromide can be prepared by the reaction of SOCl<sub>2</sub> with HBr at low temperatures. A mixed halide compound can be prepared by means of the following reaction.



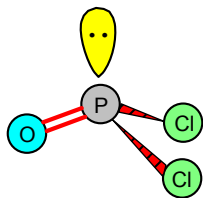
The analogous selenium compounds are obtained by the reaction of SeO<sub>2</sub> with selenium(IV) halides. For example,



Selenyl chloride is the most important of the three selenium compounds having the formula SeOX<sub>2</sub>, and it has been used as a nonaqueous solvent. The liquids have substantial conductivity so it is presumed that some autoionization occurs in both SeOCl<sub>2</sub> and SOCl<sub>2</sub>.



In the case of the thionyl and selenyl compounds the structures are pyramidal and contain only a plane of symmetry that bisects the P and O atoms and the Cl–P–Cl bond angle (*C<sub>s</sub>*):



In SOF<sub>2</sub>, the S–O bond distance is 141 pm, whereas in SOCl<sub>2</sub> it is 145 pm. A possible explanation for this is that the fluorine atoms remove more electron density from the sulfur atom resulting in a greater extent of *p*<sub>π</sub>–*d*<sub>π</sub> donation from

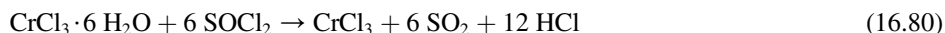


the oxygen atom. Thus, the S = O bond order is increased slightly in SOF<sub>2</sub> compared to SOCl<sub>2</sub>. Properties of thionyl and selenyl halides are summarized in [Table 16.5](#).

Chemically, thionyl chloride can function as either a Lewis acid (sulfur is the acceptor atom) or a Lewis base (oxygen is the usual electron donor atom). It behaves in the same manner as other oxyhalides of sulfur and selenium with regard to hydrolysis, and the reaction can be shown as



In fact, thionyl chloride can be used to completely dehydrate metal salts to produce anhydrous solids as illustrated by the following reaction.



Metal chlorides can also be produced from the oxides or hydroxides using SOCl<sub>2</sub>. Thionyl chloride will also react with many organic compounds. For example, the reaction with an alcohol can be shown as



### 16.10.2 Oxidation State +6

The oxidation of SO<sub>2</sub> with chlorine results in the formation of sulfuryl chloride, SO<sub>2</sub>Cl<sub>2</sub>.



This compound can be considered as the acid chloride of sulfuric acid, (HO)<sub>2</sub>SO<sub>2</sub>. As shown in the following equations it is readily hydrolyzed, and it also reacts with ammonia.



Sulfuryl fluoride is prepared by the reaction of SO<sub>2</sub> or SO<sub>2</sub>Cl<sub>2</sub> with fluorine,

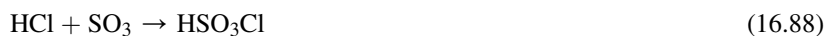


but it can also be prepared by the reaction of SF<sub>6</sub> with SO<sub>3</sub>.



The S–F bonds in SO<sub>2</sub>F<sub>2</sub> are short compared to the expected single bond distance, and they are more resistant to hydrolysis. Mixed halide compounds are obtained by the reaction of SO<sub>2</sub>F<sub>2</sub> and SO<sub>2</sub>Cl<sub>2</sub>. Selenium(VI) oxydifluoride, SeO<sub>2</sub>F<sub>2</sub>, can be prepared by heating SeO<sub>3</sub> with KBF<sub>4</sub> or by reacting BaSeO<sub>4</sub> with HSO<sub>3</sub>F.

Fluorosulfonic acid, HSO<sub>3</sub>F, and chlorosulfonic acid, HSO<sub>3</sub>Cl, are reactive compounds that have a variety of uses. Preparation of the compounds can be carried out by the following reactions.



Both the chloride and fluoride compounds hydrolyze readily to produce sulfuric acid.

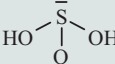
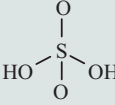
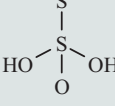
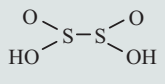
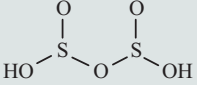
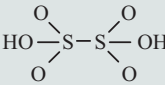
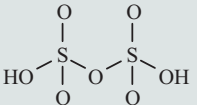
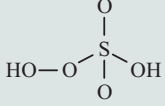
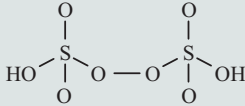
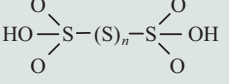


Fluorosulfonic acid is a fluorinating agent, and it also functions as an acid catalyst for alkylation, polymerization, and other reactions. Finally, both HSO<sub>3</sub>F and HSO<sub>3</sub>Cl are important sulfonating agents, reacting with organic compounds to introduce the SO<sub>3</sub>H group into various materials.

## 16.11 OXYACIDS OF SULFUR, SELENIUM, AND TELLURIUM

Sulfur, selenium, and tellurium form a variety of oxyacids. Certain of these compounds and their derivatives are of great importance with the most important of all these compounds being sulfuric acid. Because of the extraordinary significance of that compound, it will be described in more detail in a separate section. The major oxyacids of sulfur are shown in Table 16.6. Differences between the sulfur oxyacids and those of selenium and tellurium will be discussed later.

**TABLE 16.6** Major Oxyacids of Sulfur

Name	Formula	Structure
Sulfurous <sup>a</sup>	H <sub>2</sub> SO <sub>3</sub>	
Sulfuric	H <sub>2</sub> SO <sub>4</sub>	
Thiosulfuric	H <sub>2</sub> S <sub>2</sub> O <sub>3</sub>	
Dithionous <sup>a</sup>	H <sub>2</sub> S <sub>2</sub> O <sub>4</sub>	
Pyrosulfurous	H <sub>2</sub> S <sub>2</sub> O <sub>5</sub>	
Dithionic	H <sub>2</sub> S <sub>2</sub> O <sub>6</sub>	
Pyrosulfuric	H <sub>2</sub> S <sub>2</sub> O <sub>7</sub>	
Peroxymonosulfuric	H <sub>2</sub> SO <sub>5</sub>	
Peroxydisulfuric	H <sub>2</sub> S <sub>2</sub> O <sub>8</sub>	
Polythionic	H <sub>2</sub> S <sub>n+2</sub> O <sub>6</sub>	

<sup>a</sup> Only salts exist.

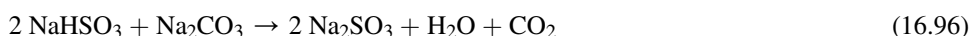
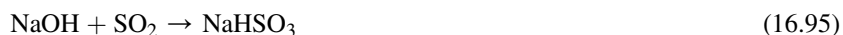
### 16.11.1 Sulfurous Acid and Sulfites

When sulfur dioxide dissolves in water, an acidic solution is formed. The gas is quite soluble (at 20 °C, 3940 cm<sup>3</sup> of SO<sub>2</sub> dissolves in 100 g of H<sub>2</sub>O at 1 atm), and the acidic solution is called “sulfurous” acid. Presumably, the reactions that take place can be written as follows.

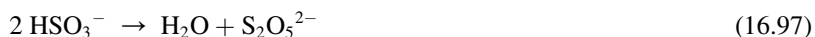


However, other species are also present (e.g., S<sub>2</sub>O<sub>5</sub><sup>2-</sup>), and most of the SO<sub>2</sub> is simply dissolved in the water without undergoing reaction. The free, undissociated acid, H<sub>2</sub>SO<sub>3</sub>, if it exists, is present in very small concentrations.

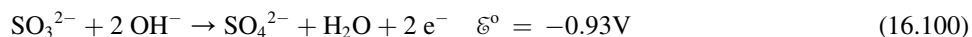
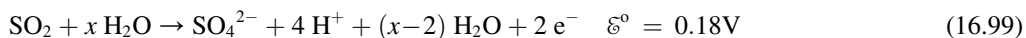
Both “normal” sulfite (containing SO<sub>3</sub><sup>2-</sup>) and acid sulfite or bisulfite (containing HSO<sub>3</sub><sup>-</sup>) salts exist with salts of the Group IA and Group IIA metals being most common. The sodium salts can be prepared as shown in the following reactions.



Bisulfite salts are not stable owing to the reactions



Chemically, sulfites and bisulfites are moderate reducing agents.



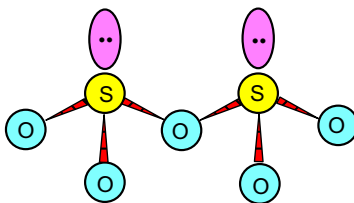
A number of oxidizing agents, among them MnO<sub>4</sub><sup>-</sup>, Cl<sub>2</sub>, I<sub>2</sub>, and Fe<sup>3+</sup>, will oxidize sulfite to sulfate.

Thiosulfates are obtained by adding sulfur to boiling solutions containing sulfites.



Thiosulfate ion is a good complexing ligand, and it forms stable complexes with silver. This reaction is the basis for removing the unreacted silver halide from photographic film.

Pyrosulfurous acid, H<sub>2</sub>S<sub>2</sub>O<sub>5</sub>, is not a stable compound. Salts of the acid are stable and they contain the S<sub>2</sub>O<sub>5</sub><sup>2-</sup> ion (sometimes considered as SO<sub>2</sub>·SO<sub>3</sub><sup>2-</sup>) that has the structure



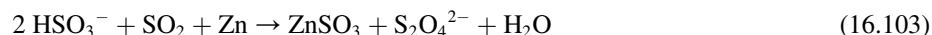
Salts of pyrosulfurous acid can be prepared by adding SO<sub>2</sub> to solutions containing sulfite salts. Dissolving a pyrosulfite in water followed by the addition of an acid produces the bisulfite ion and SO<sub>2</sub>.



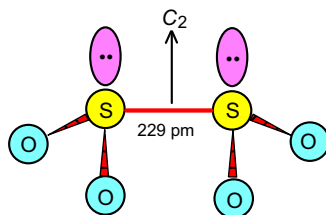
The redox chemistry of pyrosulfites is similar to that of sulfites.

### 16.11.2 Dithionous Acid and Dithionites

Dithionous acid, H<sub>2</sub>S<sub>2</sub>O<sub>4</sub>, is not obtainable as a pure compound, but several salts of the acid are known. The acid formally contains sulfur(III), and this suggests reduction of sulfites as a method of preparation. Such a process is illustrated in the following equation.



The structure of the  $\text{S}_2\text{O}_4^{2-}$  ion as it occurs in metal salts possesses one  $C_2$  axis and two mirror planes that intersect along it so the  $\text{S}_2\text{O}_4^{2-}$  ion has  $C_{2v}$  symmetry.



In this structure, the oxygen atoms are in eclipsed positions. This ion is rather unstable, and  $\text{SO}_2$  exchange occurs when labeled  $\text{SO}_2$  is added to a solution containing  $\text{S}_2\text{O}_4^{2-}$ . The long (239 pm), weak S—S bond is presumed to result from several factors. Typical S—S single bonds are about 210 pm in length. It is believed that the sulfur atoms have  $dp$  hybridization and that repulsion of the lone pairs on the sulfur atoms and repulsion of the oxygen atoms in eclipsed positions weakens the S—S bond.

In the presence of water, decomposition of  $\text{S}_2\text{O}_4^{2-}$  occurs according to the equation



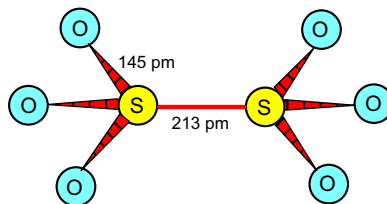
Dithionites also decompose to produce radical anions.



Air will oxidize dithionite ion (which allows it to be used as an oxygen trap), and many metals are reduced to the zero oxidation state by the ion.

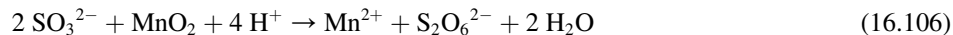
### 16.11.3 Dithionic Acid and Dithionates

The oxidation state of sulfur in  $\text{H}_2\text{S}_2\text{O}_6$  is formally +5. This acid is not obtained as a pure compound, although a sizable number of dithionate salts are known. The structure of the dithionate ion is.



A bond distance of 145 pm for S—O is comparable to that in  $\text{SO}_4^{2-}$  and is indicative of some double bond character. However, the S—S bond distance is slightly longer than that for a normal single bond.

Dithionates are prepared by the oxidation of sulfites,



or the oxidation of  $\text{SO}_2$  by  $\text{Fe}^{3+}$  in  $\text{Fe}_2(\text{SO}_3)_3$ .



At high temperatures, solid metal dithionates disproportionate to give a metal sulfate and  $\text{SO}_2$  with the general reaction being shown as follows.



Toward transition metals, the dithionate ion behaves as a bidentate ligand, and it forms a large number of complexes.

### 16.11.4 Peroxydisulfuric Acid and Peroxydisulfates

Peroxydisulfuric acid,  $\text{H}_2\text{S}_2\text{O}_8$ , is a colorless solid that has a melting point of  $65^\circ\text{C}$ . The acid and its salts are strong oxidizing agents, and the acid is not very stable. The sodium, potassium, and ammonium salts are most commonly used in oxidation reactions. The peroxydisulfate ion is generated by the anodic oxidation of bisulfate as represented by the equation



Hydrolysis of  $\text{H}_2\text{S}_2\text{O}_8$  produces  $\text{H}_2\text{O}_2$  and sulfuric acid.



The reactions shown above are the basis for a commercial preparation of hydrogen peroxide. The O—O linkage in the peroxydisulfate ion is rather unstable, so heating these solid compounds produces elemental oxygen.



Although peroxydisulfuric acid is itself of little importance except as an intermediate in the preparation of  $\text{H}_2\text{O}_2$ , the salts are important oxidizing agents for synthetic purposes.

Peroxymonosulfuric acid,  $\text{H}_2\text{S}_2\text{O}_5$ , is obtained by the partial hydrolysis of  $\text{H}_2\text{S}_2\text{O}_8$ .



It can also be prepared by the reaction of  $\text{H}_2\text{O}_2$  and chlorosulfonic acid.



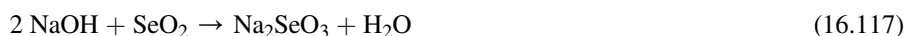
Peroxymonosulfuric acid, sometimes called *Caro's acid*, and its salts are of much less importance than the peroxydisulfates. Although polythionic acids having the general formula  $\text{H}_2\text{S}_n\text{O}_6$  and salts of those acids are known, they are of little importance and will not be discussed further.

### 16.11.5 Oxyacids of Selenium and Tellurium

Selenous and tellurous acids can be obtained by the hydrolysis of the tetrahalides,



or by slow evaporation of solutions of the dioxides in water. Unlike  $\text{H}_2\text{SO}_3$ , solid  $\text{H}_2\text{SeO}_3$  can be obtained. A large number of selenites and tellurites are known, and typical preparations include the following reactions.



In either concentrated solutions or the molten acid, selenite is partially converted by dehydration to pyroselenite,  $\text{Se}_2\text{O}_5^{2-}$ .



Tellurites also form a series of polytellurites ( $\text{Te}_4\text{O}_9^{2-}$ ,  $\text{Te}_6\text{O}_{13}^{2-}$ , etc.) that can be obtained as solid salts.

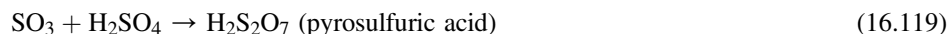
Selenic and telluric acids contain the central atom in the +6 oxidation state, but they are quite different. Selenic acid,  $\text{H}_2\text{SeO}_4$ , behaves very much like  $\text{H}_2\text{SO}_4$  in most of its chemical properties. Most selenates are isomorphous with the corresponding sulfates, and some pyroselenates are known. On the other hand, telluric acid has the formula  $\text{H}_6\text{TeO}_6$  or  $\text{Te}(\text{OH})_6$ . Although this formula is equivalent to  $\text{H}_2\text{TeO}_4 \cdot 2\text{H}_2\text{O}$ , the molecule is octahedral and is not a dihydrate. It is isoelectronic with the hydroxocomplexes  $\text{Sn}(\text{OH})_6^{2-}$  and  $\text{Sb}(\text{OH})_6^-$ . As a result of its structure, telluric acid is a weak acid (see Chapter 6). Several types of salts of telluric acid are known in which one or more of the hydrogen atoms is replaced (e.g.,  $\text{KH}_5\text{TeO}_6$  or  $\text{K}[\text{TeO}(\text{OH})_5]$ ,  $\text{Hg}_3\text{TeO}_6$ , and  $\text{Li}_2\text{H}_4\text{TeO}_6$ ). Neither selenic nor telluric acid approaches the commercial importance of sulfuric acid. But then, no compound does.

## 16.12 SULFURIC ACID

It is hard to overemphasize the importance of sulfuric acid in chemistry and in the economy. As far as chemicals are concerned, its impact on the economy is second to none. Without sulfuric acid utilization on a large scale, many of the products in general use would not exist, and our very way of life would be altered. The next section will present some of the essential chemistry of  $\text{H}_2\text{SO}_4$  to show why it is so important.

### 16.12.1 Preparation of Sulfuric Acid

In the 1800s, the most important method for producing  $\text{H}_2\text{SO}_4$  was by the *lead chamber process*. Today, sulfuric acid is produced by a method known as the *contact process*. In the contact process, sulfur is burned to give  $\text{SO}_2$  or the required  $\text{SO}_2$  is recovered from coal burning or ore roasting processes. The  $\text{SO}_2$  is then oxidized in the presence of a catalyst to produce  $\text{SO}_3$  (see Section 16.7.2). Typical catalysts are spongy platinum or sodium vanadate. Next, the  $\text{SO}_3$  is dissolved in 98% sulfuric acid.



Then, water can be added to produce any desired concentration of sulfuric acid, even 100%.



The preparation of  $\text{SO}_2$  from burning fossil fuels and roasting ores have been described earlier in this chapter as sources of  $\text{SO}_2$  that are replacing the burning of sulfur.

In the older lead chamber process,  $\text{SO}_2$ , oxygen, steam,  $\text{NO}$  and  $\text{NO}_2$  were introduced into lead-lined chambers. The nitrogen oxides catalyze the oxidation of  $\text{SO}_2$ , and the reactions that occur can be summarized as follows.



The actual reactions are quite complex, however, and they will not be described further because the process is no longer of commercial importance.

### 16.12.2 Physical Properties of Sulfuric Acid

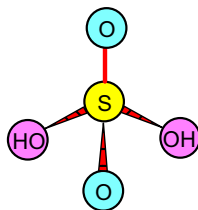
The usual commercial form of  $\text{H}_2\text{SO}_4$  is the concentrated acid that contains about 98%  $\text{H}_2\text{SO}_4$  (18 M). It is a thick oily dense liquid ( $1.85 \text{ g cm}^{-3}$  at  $25^\circ\text{C}$ ) having a high viscosity resulting from its extensive association by hydrogen bonding. The heat of dilution of 98%  $\text{H}_2\text{SO}_4$  is very high (about  $-879 \text{ kJ mol}^{-1}$  at infinite dilution), and this can cause boiling of the water as it is added in the dilution process so that spattering and splashing of the mixture can occur. Dilution of concentrated  $\text{H}_2\text{SO}_4$  should be carried out by adding the acid slowly to water while stirring the mixture constantly. The concentrated acid has a very low vapor pressure and is a strong dehydrating agent. Also, a number of well-defined hydrates are known, e.g.,  $\text{H}_2\text{SO}_4 \cdot \text{H}_2\text{O}$  (m.p.  $8.5^\circ\text{C}$ ),  $\text{H}_2\text{SO}_4 \cdot 2 \text{H}_2\text{O}$  (m.p.  $-39.5^\circ\text{C}$ ), and  $\text{H}_2\text{SO}_4 \cdot 4 \text{H}_2\text{O}$  (m.p.  $-28.2^\circ\text{C}$ ). Some of the physical properties of  $\text{H}_2\text{SO}_4$  are shown below.

Melting point	$10.4^\circ\text{C}$
Boiling point	$290^\circ\text{C}$ (with decomp.)
Dielectric constant	100
Density ( $25^\circ\text{C}$ )	$1.85 \text{ g cm}^{-3}$
Viscosity ( $25^\circ\text{C}$ )	24.54 cp

Sulfuric acid in which additional  $\text{SO}_3$  is dissolved behaves as though it were greater than 100% sulfuric acid because the solution can be diluted with water and still be 100% sulfuric acid. Such a solution of  $\text{SO}_3$  in  $\text{H}_2\text{SO}_4$  (written as  $\text{H}_2\text{S}_2\text{O}_7$ ) is called *oleum*. The concentration of oleum is specified in terms of the percent of free  $\text{SO}_3$  present

and typical commercial oleum ranges from 10% to 70% free  $\text{SO}_3$ . Oleum undergoes reactions similar to the concentrated acid.

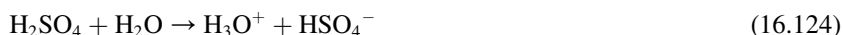
The structure of the  $\text{H}_2\text{SO}_4$  molecule is



with the sulfur atom surrounded by a tetrahedral arrangement of oxygen atoms with S–O bond distances of 151 pm. This bond distance is indicative of some multiple bonding between S and O atoms of the  $p\pi-d\pi$  type. In view of the fact that the S atom has empty  $3d$  orbitals only slightly above the valence shell in energy, it is plausible to expect some donation of electron density from the filled  $p$  orbitals on the oxygen atoms to empty  $d$  orbitals on the sulfur atom.

### 16.12.3 Chemical Properties of Sulfuric Acid

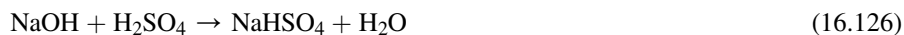
The dominant chemical property of  $\text{H}_2\text{SO}_4$  is its strong acidity. Dissociation in the first step,



is complete in dilute solutions. For the second step,



$K_{a2} = 1.29 \times 10^{-2}$  at 18 °C. Therefore, in the second step, sulfuric acid is by no means a very strong acid at reasonable concentrations. Salts of  $\text{H}_2\text{SO}_4$  consist of both the normal sulfates containing  $\text{SO}_4^{2-}$  and acid sulfates or bisulfates containing  $\text{HSO}_4^-$ . The bisulfates can be obtained by partial neutralization of the acid,



or by heating other salts with  $\text{H}_2\text{SO}_4$ .



When bisulfates are dissolved in water, the solutions are somewhat acidic because of the dissociation of the  $\text{HSO}_4^-$  ion as shown in Eq. (16.125).

Normal sulfates result from several processes including the following types of reactions.

(a) complete neutralization of sulfuric acid:



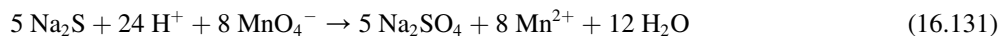
(b) dissolution of a metal in sulfuric acid:



(c) metathesis:

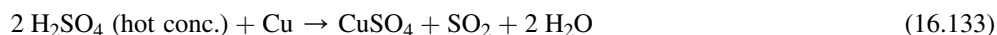


(d) oxidation reactions:



Although the reaction of a sulfide is shown above, sulfites and other sulfur-containing compounds can be oxidized to sulfates by using the proper oxidizing agents under the appropriate conditions.

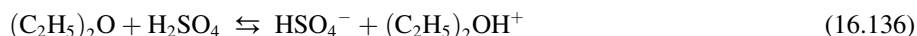
Most sulfates except those of  $\text{Ca}^{2+}$ ,  $\text{Ba}^{2+}$ ,  $\text{Pb}^{2+}$ ,  $\text{Sr}^{2+}$ , and  $\text{Eu}^{2+}$  are soluble in water. Although  $\text{H}_2\text{SO}_4$  does not normally react as an oxidizing agent, the hot, concentrated acid does behave in that way.



In addition to being acidic in aqueous solutions, bisulfate salts, such as  $\text{NaHSO}_4$ , lose water on heating strongly to produce pyrosulfates.



Sulfuric acid has received considerable study as a nonaqueous solvent (see Chapter 6). It is, of course, a strongly acidic solvent, and it has a  $K_f$  value of  $-6.15^\circ/\text{molal}$ . In 100%  $\text{H}_2\text{SO}_4$ , conductivity measurements and cryoscopic studies show that protonation of many substances occurs even though they are not normally bases in the usual sense. For example, organic compounds such as acetic acid and ether function as proton acceptors.



When behaving as a nonaqueous solvent, it is readily apparent that substances that are slightly acidic in water may behave as bases in liquid  $\text{H}_2\text{SO}_4$ . Of course, any substance that reacts as a base in water becomes a much stronger base toward  $\text{H}_2\text{SO}_4$  because it is such a strong acid. It is believed that 100%  $\text{H}_2\text{SO}_4$  also undergoes some autoionization.

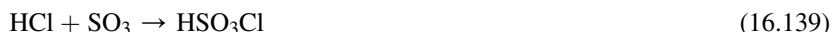


The reaction of sulfuric acid with nitric acid generates the nitronium ion,  $\text{NO}_2^+$ .



The  $\text{NO}_2^+$  ion is the attacking species involved in nitration reactions so  $\text{H}_2\text{SO}_4$  serves to increase the concentration of a positive attacking species and is, therefore, functioning as an acid catalyst (see Chapter 6).

Many derivatives of sulfuric acid are also important. In particular, sulfonates are used in the preparation of detergents such as  $\text{CH}_3(\text{CH}_2)_{11}\text{C}_6\text{H}_4\text{SO}_3^- \text{Na}^+$ . Chlorosulfonic acid,  $\text{HSO}_3\text{Cl}$ , can be prepared by the reaction of  $\text{HCl}$  with  $\text{SO}_3$ .



This compound can function as an acid, but it also contains a highly reactive S—Cl bond that will react with water, alcohols, etc.

Sulfuryl chloride,  $\text{SO}_2\text{Cl}_2$ , is a highly reactive and useful intermediate that can be prepared by the reaction of  $\text{SO}_2$  and  $\text{Cl}_2$ .



Other derivatives and some of the chemical uses of sulfuric acid are described in other sections.

#### 16.12.4 Uses of Sulfuric Acid

Sulfuric acid is used in the largest quantity of any chemical. Approximately 79 billion pounds were produced in 2014. Although many of the uses of sulfuric acid now are the same as they have been for many years, there has been a great change in some of the uses. These changes reflect changes in our society and economy. For example, in 1959, only 32% of the sulfuric acid produced was used in the manufacture of fertilizer. By 1977, that use accounted for 65% of the sulfuric acid, and the total amount of acid used was then approximately twice that consumed in 1959. About 11% of the sulfuric acid was used in petroleum refining in 1959. By 1977, that use accounted for about 5% of the  $\text{H}_2\text{SO}_4$  produced. Also, about 25% of the sulfuric acid produced in 1959 was used in the manufacture of other chemicals, but in 1977 that use amounted to 5% of the  $\text{H}_2\text{SO}_4$  manufactured. In the period from 1959 to 1977, annual production of sulfuric acid rose from about 36 billion pounds to about 68 billion pounds. For the last several years, the production



has amounted to about 80–82 billion pounds. The percentage of H<sub>2</sub>SO<sub>4</sub> in each use is approximately the same now as it was in 1977.

In the manufacture of fertilizers, sulfuric acid is used to digest phosphate rock (largely Ca<sub>3</sub>(PO<sub>4</sub>)<sub>2</sub> and Ca<sub>5</sub>(PO<sub>4</sub>)<sub>3</sub>F) to make phosphoric acid or the calcium salts such as Ca(H<sub>2</sub>PO<sub>4</sub>)<sub>2</sub>. Some of the phosphoric acid is also used to make ammonium phosphate. More details on the manufacture of fertilizer were presented in Chapter 14.

In refining of petroleum, sulfuric acid is used as a catalyst for alkylation reactions and in the manufacture of other organic derivatives. Chemicals production includes virtually every heavy chemical industry. For example, the production of sodium and aluminum sulfates, hydrochloric and hydrofluoric acids, insecticides, detergents, and many other chemicals all involve the use of H<sub>2</sub>SO<sub>4</sub>.

Sulfuric acid is used in large quantities in the “pickling” of steel to remove oxide coatings by dissolving them. Another large use of sulfuric acid involves its reaction with ammonia obtained in the conversion of coal to coke.



In that process, a large quantity of ammonium sulfate is produced, and it is used in fertilizers because it is a relatively inexpensive source of nitrogen. Finally, a large quantity of sulfuric acid is used as the electrolyte in millions of automobile batteries. With so many processes that consume enormous quantities of H<sub>2</sub>SO<sub>4</sub>, it not surprising that the production and consumption of sulfuric acid provides a “barometer” to the general health of the economy!

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## PROBLEMS

1. Explain why the addition of two electrons to S(g) is less endothermic than the addition of two electrons to O(g).
2. Write the equation for the reaction of SF<sub>4</sub> with water. Why does SF<sub>4</sub> react so differently when compared to SF<sub>6</sub>?
3. Provide an explanation of why SF<sub>6</sub> sublimates at –63.8 °C but SF<sub>4</sub> boils at –38 °C.
4. Explain using structures how SF<sub>4</sub> can behave as both a Lewis acid and a Lewis base.
5. Draw the structure of each of the following molecules.
  - (a) SeCl<sub>4</sub>
  - (b) Peroxydisulfuric acid
  - (c) Dithionic acid
  - (d) SF<sub>4</sub>
  - (e) SO<sub>3</sub>
6. Discuss briefly why, although SF<sub>6</sub> is chemically inert, TeF<sub>6</sub> is more reactive.
7. Draw the structure of each of the following molecules.
  - (a) Dithionate ion
  - (b) Thionyl chloride
  - (c) S<sub>2</sub>Cl<sub>2</sub>
  - (d) H<sub>2</sub>S<sub>6</sub>
  - (e) Sulfuryl chloride

8. Provide an explanation of why  $\text{H}_2\text{Se}$  and  $\text{H}_2\text{Te}$  are less stable than  $\text{H}_2\text{S}$ .
9. Write equations to show the preparation of:
- $\text{H}_2\text{Te}$
  - $\text{SO}_2\text{Cl}_2$
10. Write balanced equations for the following processes.
- Preparation of hydrogen telluride (starting with tellurium)
  - Decomposition of potassium peroxydisulfate
  - Preparation of sulfuryl chloride (starting with the elements)
  - Preparation of sodium thiosulfate
  - Preparation of  $\text{Na}_2\text{S}_2\text{O}_7$
11. Complete and balance the following.
- $\text{SO}_2\text{Cl}_2 + \text{C}_2\text{H}_5\text{OH} \rightarrow$
  - $\text{BaO} + \text{SeO}_3 \rightarrow$
  - $\text{NaOH} + \text{SO}_2 \rightarrow$
  - $\text{HOSO}_2\text{Cl} + \text{H}_2\text{O}_2 \rightarrow$
  - $\text{PbS} + \text{HCl} \rightarrow$
12. Complete and balance the following.
- $\text{Cu} + \text{H}_2\text{SO}_4 \rightarrow$
  - $\text{Sb}_2\text{S}_3 + \text{HCl} \rightarrow$
  - $\text{Na}_2\text{SO}_3 + \text{S} \rightarrow$
  - $\text{H}_2\text{S}_2\text{O}_7 + \text{H}_2\text{O} \rightarrow$
  - $\text{HNO}_3 + \text{Se} \rightarrow$
13. Complete and balance the following.
- $\text{HNO}_3 + \text{S}_8 \rightarrow$
  - $\text{S}_2\text{Cl}_2 + \text{NH}_3 \rightarrow$
  - $\text{S}_4\text{N}_4 + \text{Cl}_2 \rightarrow$
  - $\text{SO}_2 + \text{PCl}_5 \rightarrow$
  - $\text{CaS}_2\text{O}_6 \rightarrow$
14. When  $\text{S}_4\text{N}_4$  reacts to add chlorine atoms, where would you expect them to be bonded? Why?
15. For each of the following species, draw the structure, determine the symmetry elements it possesses, and determine the symmetry type.
- $\text{OSCl}_2$
  - $\text{SeCl}_2$
  - $\text{S}_2\text{O}_3^{2-}$
  - $\text{S}_8$
  - $\text{SO}_2\text{Cl}_2$

# Halogens

The elements in Group VIIA of the periodic table are called the *halogens*. This name comes from two Greek words, *halos*, meaning “salt” and *genes*, meaning “born.” Thus, the halogens are the “salt formers,” which they are in fact as well as in name. Because of the reactivity of these elements, they do not occur free in nature. Thus, although compounds of these elements (particularly salt) were known to the ancients, the elements themselves have been known for a much shorter time. Chlorine was prepared in 1774 by Scheele by the reaction of HCl with MnO<sub>2</sub>. Davy suggested the name based on the Greek *chloros*, meaning “greenish yellow.” Bromine was discovered in 1826 by Balard. He eventually adapted the name from the Greek word *bromos*, meaning “stench.” The name iodine is adapted from the Greek word *iodes*, meaning “violet.” It was discovered about 1812 by Courtois in his studies on kelp. Early work on iodine was also done by Gay-Lussac in 1813–1814. Although heating *fluorspar* with sulfuric acid was known to produce an acid that etches glass as long ago as 300 years, elemental fluorine was produced much later. This was due to the difficulty in obtaining pure HF and the fact that an electrochemical process had to be used to obtain elemental fluorine. This was finally done in 1886 by Moissan. The name fluorine comes from the Latin word *fluere* which means “to flow.” Astatine was first reported in 1940, and all of its isotopes are radioactive. Its name is derived from a Greek word meaning “unstable.” It is available only as a product of nuclear reactions and, consequently, we will deal very briefly with its chemistry.

## 17.1 OCCURRENCE

Fluorine occurs in nature in the form of the minerals *fluorite*, CaF<sub>2</sub>, *cryolite*, Na<sub>3</sub>AlF<sub>6</sub>, and *fluoroapatite*, Ca<sub>5</sub>(PO<sub>4</sub>)<sub>3</sub>F, and one commercial source of natural cryolite is Greenland. Both of the other minerals are widespread in nature, although the major use of fluoroapatite is in the production of fertilizers, not as a primary source of fluorine. Extensive fluorite deposits are found in Southeastern Illinois and Northwestern Kentucky. From the standpoint of fluorine utilization, both cryolite and fluorite are extremely important minerals. Cryolite is used as the electrolyte in the electrochemical production of aluminum from bauxite, and fluorite is used as a flux in making steel. Today, most of the cryolite used is synthetic rather than the naturally occurring mineral.

Chlorine is abundantly available in NaCl and in salt water. Hence, the quantity of chlorine combined in these natural sources is enormous. The Great Salt Lake contains 23% salt and the Dead Sea contains about 30%. Chlorine also occurs in a few minerals, but the abundance of naturally occurring salt water makes these of little importance.

Bromine is found as the bromide salts of Group IA and Group IIA metals, usually along with the chlorides. Bromine and iodine are also found in compounds contained in brines and seawater. A few minerals (e.g., *lautorite*, Ca(IO<sub>3</sub>)<sub>2</sub>) are also known that contain iodine, and it is found in plants and animals from the sea.

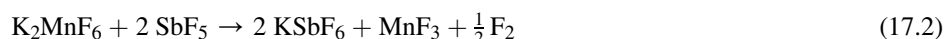
## 17.2 THE ELEMENTS

Fluorine is produced by the electrolysis of a molten mixture of KF and HF. Fluorides readily form bifluorides (hydrogen difluorides) by the reaction



so the addition of HF to KF produces KHF<sub>2</sub> (m.p. 240 °C). This melt will dissolve additional HF until a composition in the range of KF·2 HF to KF·3 HF is obtained. This mixture melts at 80–90 °C, and it is electrolyzed in containers made of Monel metal, copper, steel, and other materials that resist attack by F<sub>2</sub> and HF by forming a protective fluoride coating.

Fluorine has been prepared chemically in recent years by means of the reaction



The basis for this reaction is the fact that the strong Lewis acid, SbF<sub>5</sub>, removes fluoride ions from MnF<sub>6</sub><sup>2-</sup> to produce MnF<sub>4</sub> which is thermodynamically unstable and decomposes to produce MnF<sub>3</sub> and F<sub>2</sub>.

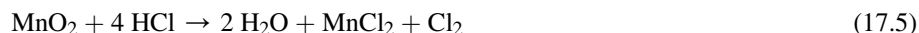
Chlorine is produced in largest quantity by the electrolysis of NaCl solutions in a process represented by the reaction



Actually, the equation above represents the process in which a diaphragm is used to keep the anode and cathode compartments separate. This process is of enormous industrial importance because it produces chlorine, sodium hydroxide, and hydrogen, all of which are used in large quantities. In another process, a mercury cathode is employed and sodium amalgam is produced as shown in the equation



As a laboratory preparative procedure, the oxidation of  $\text{Cl}^-$  by  $\text{MnO}_2$  first employed by Scheele is still used,



but oxidizing agents other than  $\text{MnO}_2$  can be used. Another method that has also been used commercially is the oxidation of HCl.



Oxides of nitrogen catalyze this process and removal of  $\text{H}_2\text{O}$  favors the reaction. Some chlorine is also produced in the electrolysis of molten NaCl which is the chief preparation of elemental sodium.



Bromine is obtained from seawater by oxidation of  $\text{Br}^-$  with  $\text{Cl}_2$  followed by sweeping out the  $\text{Br}_2$  with a stream of air.



The oxidation of HBr by  $\text{O}_2$  or other oxidizing agents can also be utilized.

Iodine is obtained by oxidizing iodides from seawater or brines using  $\text{Cl}_2$ , concentrated  $\text{H}_2\text{SO}_4$ ,  $\text{Fe}^{3+}$ , or other oxidizing agents. Astatine is produced naturally by the radioactive decay of uranium or thorium. Production of  $^{211}\text{At}$  is also accomplished by bombarding  $^{209}\text{Bi}$  with alpha particles,



Some of the properties of the halogens are summarized in [Table 17.1](#).

Naturally occurring fluorine consists of 100% of the isotope 19, chlorine consists of about 75% of isotope 35 and about 25% of isotope 37, bromine consists of about equal abundances of isotopes 79 and 81, and iodine consists of 100% of isotope 127. Chemically, the halogens are all oxidizing agents with the strength decreasing as  $\text{F}_2 > \text{Cl}_2 > \text{Br}_2 > \text{I}_2$ . The elements oxidize most of the other elements to produce a variety of ionic and covalent halides. The halogens, except for  $\text{F}_2$ , dissolve in water to undergo a disproportionation reaction that can be represented as follows:



**TABLE 17.1** Properties of the Halogens

	$\text{F}_2$	$\text{Cl}_2$	$\text{Br}_2$	$\text{I}_2$
m.p., °C	-220	-101	-7.25	113.6
b.p., °C	-188	-34.1	59.4	185
X-X bond energy, $\text{kJ mol}^{-1}$	153	239	190	149
X-X distance, pm	142	198	227	267
Electronegativity (Pauling)	4.0	3.0	2.8	2.5
Electron affinity, $\text{kJ mol}^{-1}$	339	355	331	302
Single bond radius, pm	71	99	114	133
Anion ( $\text{X}^-$ ) radius, pm	133	181	196	220

However, as a result of its extreme reactivity, fluorine reacts with water to liberate oxygen.



## 17.3 INTERHALOGENS

Halogens form not only salts but they also form many covalent compounds including those composed of two different halogens. These are known as the *interhalogens*. The general formula for these compounds is  $\text{XX}'_n$  where  $X'$  is the lighter halogen (having higher electronegativity) and  $n$  is an odd number. Although there is potentially a sizeable number of such compounds, a much smaller number of interhalogens actually have been identified. For example, the formula where  $n = 7$  occurs only for the compound  $\text{IF}_7$  and  $n = 5$  only for  $\text{ClF}_5$ ,  $\text{BrF}_5$ , and  $\text{IF}_5$ .

The electronic configuration of the halogens,  $ns^2 np^5$ , indicates that these atoms should readily form only one covalent bond. Formation of additional covalent bonds can be accomplished only when unshared pairs of electrons become unpaired or the valence shell expands to hold more than eight electrons. Unpairing two electrons means that two additional covalent bonds to other halogens atoms can form. Therefore, the total number of bonds to the central (heavier) halogen atom in interhalogen molecules is 1, 3, 5, or 7. The compounds contain no unpaired electrons so they are diamagnetic. Most of the known interhalogens are fluorides in accord with the principle that the greater electronegativity difference leads to greater bond polarity that results in greater stability. Also, as the size of the central halogen increases, a greater number of lighter halogens can be held around it. For example, placing more than five fluorine atoms around a chlorine or bromine atom is unlikely, but iodine can bond to seven fluorine atoms. Of course, it is also easier to oxidize the iodine atom to a +7 oxidation state than it is to oxidize either chlorine or bromine to the +7 oxidation state. Table 17.2 shows melting and boiling points for the interhalogens.

### 17.3.1 Type $\text{XX}'$

Chlorine monofluoride can be prepared by direct combination of the elements at elevated temperatures.



**TABLE 17.2** The Interhalogens

	Formula	m.p., °C	b.p., °C
Type $\text{XX}'$	ClF	-156	-100
	BrF	-33	20
	IF	—	—
	BrCl	-66	5
	ICl	27	97
	IBr	41	116
Type $\text{XX}'_3$	ClF <sub>3</sub>	-83	12
	BrF <sub>3</sub>	8	127
	IF <sub>3</sub>	—	—
	ICl <sub>3</sub>	101 (16 atm)	—
Type $\text{XX}'_5$	ClF <sub>5</sub>	-103	-14
	BrF <sub>5</sub>	-60	41
	IF <sub>5</sub>	10	101
Type $\text{XX}'_7$	IF <sub>7</sub>	6.45 (tr. pt.)	—

Bromine monofluoride is much less stable than ClF although it has been prepared by reaction of the elements diluted with nitrogen.



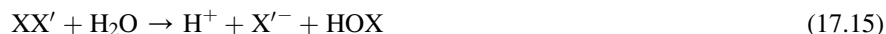
The compound spontaneously decomposes to give Br<sub>2</sub> and BrF<sub>3</sub> or BrF<sub>5</sub>. Iodine monofluoride is so unstable that its existence is somewhat questionable.

Bromine monochloride is also prepared by the combination of the elements.



It is an unstable compound but it does exist in equilibrium with the elements. Iodine monochloride is much more stable, and it is prepared by the reaction of liquid chlorine with solid iodine in stoichiometric quantities. Two forms of ICl are known: a reddish brown solid and the more stable ruby red needles that form on prolonged standing. Iodine monobromide is also prepared by combination of the elements.

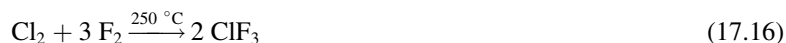
All the compounds of the type XX' are reactive materials that undergo some reactions that are similar to those of the free halogens. For example, they are all oxidizing agents and usually add to double bonds in organic molecules. They also react with water.



In these reactions, the less electronegative halogen is found in the corresponding hypohalous acid where it has a +1 oxidation state.

### 17.3.2 Type XX'<sub>3</sub>

Compounds of the type XX'<sub>3</sub> include ClF<sub>3</sub>, BrF<sub>3</sub>, and ICl<sub>3</sub> with the existence of IF<sub>3</sub> being somewhat questionable. These compounds are prepared either by direct combination of the elements or by addition of the lighter halogen to the monohalo compounds. Chlorine trifluoride is commercially available and is a widely used fluorinating agent. One of the attractive features of ClF<sub>3</sub> in this application is that it reacts with less vigor than does elemental fluorine so many fluorination reactions are less vigorous. The reaction of chlorine with an excess of fluorine yields ClF<sub>3</sub>.



Although this compound is very reactive as a fluorinating agent, it can be handled in glass or in containers made of some metals. Many organic compounds react with ClF<sub>3</sub> so vigorously that combustion occurs. The compound also reacts with bromine and iodine to yield BrF<sub>3</sub> and IF<sub>5</sub>, respectively,



Bromine trifluoride can be prepared by the reaction of Br<sub>2</sub> with F<sub>2</sub>,



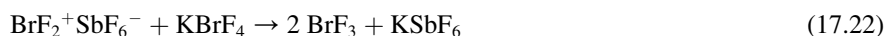
Some BrF is also formed in the reaction, and it can be separated by fractional distillation. Bromine trifluoride is a liquid at room temperature and shows an electrical conductivity high enough to indicate some autoionization which is usually represented by the reaction



The use of BrF<sub>3</sub> as a nonaqueous solvent has been the subject of a considerable amount of study. In this solvent, BrF<sub>2</sub><sup>+</sup> is the acidic species and BrF<sub>4</sub><sup>-</sup> is the basic species. Therefore, SbF<sub>5</sub> is an acid in liquid BrF<sub>3</sub> because it increases the concentration of the acidic species, BrF<sub>2</sub><sup>+</sup>, as a result of the reaction



A neutralization reaction in this solvent can be considered as a reaction of BrF<sub>2</sub><sup>+</sup> with BrF<sub>4</sub><sup>-</sup> to produce the unionized solvent. An example of this behavior is the following:



Iodine trichloride can be prepared by the reaction of an excess of liquid chlorine with solid iodine and then allowing the excess chlorine to evaporate from the solid  $\text{ICl}_3$ .

### 17.3.3 Type $\text{XX}'_5$

Chlorine pentafluoride is less stable than  $\text{ClF}_3$ , and it decomposes to  $\text{ClF}_3$  and  $\text{F}_2$  at temperatures above  $165^\circ\text{C}$ . Below that temperature, the equilibrium shown as



favors the formation of  $\text{ClF}_5$ . Bromine pentafluoride is prepared by the reaction of fluorine with  $\text{BrF}_3$  or the direct combination of bromine and fluorine. The compound is very reactive, and it is a strong fluorinating agent that reacts explosively with many organic compounds.

Iodine pentafluoride is prepared directly from the elements, and it is somewhat less reactive than  $\text{BrF}_5$ . Pure  $\text{IF}_5$  apparently undergoes some autoionization,



Above  $500^\circ\text{C}$ , the compound disproportionates to give  $\text{I}_2$  and  $\text{IF}_7$ . At the boiling point ( $101^\circ\text{C}$ ),  $\text{IF}_5$  reacts with  $\text{KF}$  to produce  $\text{IF}_6^-$ ,



Because the oxidation state of iodine in  $\text{IF}_5$  is +5, the reaction with water produces the acid containing iodine in that oxidation state,  $\text{HIO}_3$ .



### 17.3.4 Type $\text{XX}'_7$

The only example of this type of compound is  $\text{IF}_7$ . This is the result of several factors, among them the large size of the iodine atom and the easier oxidation of iodine to a +7 oxidation state. Also, the combination of iodine with fluorine represents the largest electronegativity difference possible for interhalogens. The compound is prepared by the reaction of fluorine with  $\text{IF}_5$  at elevated temperatures.



Iodine heptafluoride is a strong fluorinating agent. It reacts with water to produce the acid containing iodine in the +7 oxidation state,  $\text{HIO}_4$  (or  $\text{H}_5\text{IO}_6$ ).



### 17.3.5 Structures

The structures of the interhalogens having the formula  $\text{XX}'$  are polar diatomic molecules. In the molecules having the formula  $\text{XX}'_3$ , the central halogen has a total of 10 electrons surrounding it (7 from its own valence shell and 1 from each of the other atoms attached to it). These 10 electrons are arranged in 3 bonding pairs and 2 nonbonding pairs. This results in a structure ( $C_{2v}$ ) that can be considered as arising from a trigonal bipyramid with the two unshared pairs in equatorial positions as shown for  $\text{BrF}_3$  in Figure 17.1(a).

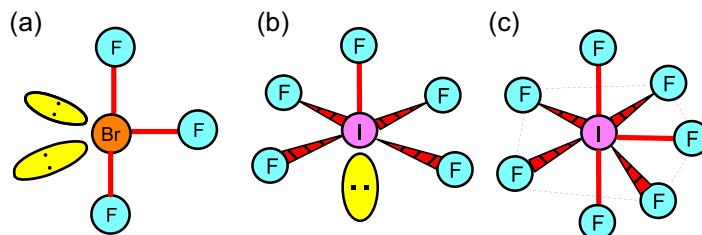
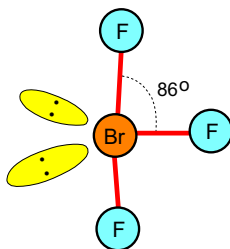


FIGURE 17.1 The structures of some interhalogens of the  $\text{XX}'_3$ ,  $\text{XX}'_5$ , and  $\text{XX}'_7$  types.

In this molecule, the repulsion between the nonbonding pairs and the bonding pairs causes the structure to be somewhat irregular in accord with the valence shell electron pair repulsion (VSEPR) prediction. Consequently, the actual structure of the molecule is



In compounds having the formula  $XX'_5$ , such as  $IF_5$ , there are 12 electrons around the central atom. Orientation of six pairs of electrons in six orbitals around a central atom usually has the orbitals directed toward the corners of an octahedron. As shown in Figure 17.1(b), one of the six positions on the octahedron is occupied by a nonbonding pair of electrons. The repulsion between the nonbonding pair of electrons and the bonding pairs causes a slight deviation from a regular structure so that the iodine atom lies slightly below the plane of the four fluorine atoms in a  $C_{4v}$  structure. Consequently, the  $X-F-X$  bond angle is about  $82^\circ$  in  $IF_5$  and about  $84^\circ$  in  $BrF_5$ . In the  $IF_5$  molecule, the axial distance,  $I-F_{ax}$ , is 184 pm whereas the equatorial distance,  $I-F_{eq}$ , is 187 pm. In the case of  $BrF_5$ , the corresponding bond distances are 169 and 177 pm, respectively. The structure of  $ClF_5$  is similar to that of  $IF_5$  and  $BrF_5$ , but the dimensions have not been as accurately determined.

In the only compound of the  $XX'_7$  type,  $IF_7$ , there are seven pairs of electrons around the iodine atom, all of which are bonding pairs. Because there is no valence shell repulsion by nonbonding pairs, the structure is that of a regular pentagonal bipyramid,  $D_{5h}$ . The  $I-F_{ax}$  distance is 179 pm whereas the  $I-F_{eq}$  distance is 186 pm.

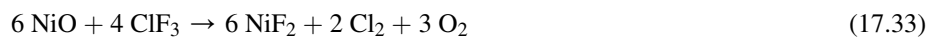
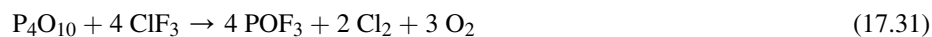
### 17.3.6 Chemical Properties

Probably the most useful interhalogens are the halogen fluorides, especially  $ClF_3$ ,  $BrF_3$ , and  $BrF_5$ . All of these compounds are very strong oxidizing agents giving reactions similar to those of fluorine. Accordingly, they must be handled with extreme caution. Because of the oxidizing strength of fluorine and these interhalogens, many materials, such as wood, concrete, or flesh, will readily burn in the presence of such oxidants. Containers made of such materials as nickel, Monel<sup>®</sup> (an alloy of nickel and copper with minor amounts of other metals), mild steel, and copper form a fluoride coating and resist further attack. All of these interhalogens are used as fluorinating agents.

Many organic compounds react vigorously or explosively with  $ClF_3$ ,  $BrF_3$ , or  $BrF_5$ . To moderate the reactions, diluents such as HF,  $CCl_4$ , or  $N_2$  are used. With the exception of the noble gases,  $N_2$ ,  $O_2$ , and  $F_2$ , all of the nonmetallic elements react with halogen fluorides to form the corresponding fluorides. The actual distribution of products depends on the reaction conditions, but, in general, these strong oxidizing agents convert the elements to fluorine compounds in the highest oxidation state possible for the elements. The following reactions of  $ClF_3$  are typical:



Metal and nonmetal oxides are converted to fluorides or oxyfluorides by reactions with halogen fluorides as illustrated by the following equations:



Metal salts are generally converted to the metal fluorides by reactions with the halogen fluorides, with the metal being oxidized to its highest oxidation state if an excess of the interhalogen is present.



Finally, the following reaction of nitrosyl fluoride, NOF, with ClF is noteworthy:



The product, NOClF<sub>2</sub>, is essentially ionic, and it contains the NO<sup>+</sup> (nitrosyl) cation and the ClF<sub>2</sub><sup>-</sup> anion. This polyhalide anion is similar to several others that are known. For example, I<sub>3</sub><sup>-</sup> results when I<sub>2</sub> is dissolved in a solution that contains KI.

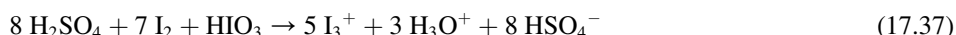


The X<sub>3</sub><sup>-</sup> and X<sup>-</sup>X<sub>2</sub><sup>-</sup> species are linear. We shall explore the characteristics and behavior of the polyhalide species in greater detail in the next section.

## 17.4 POLYATOMIC CATIONS AND ANIONS

### 17.4.1 Polyatomic Halogen Cations

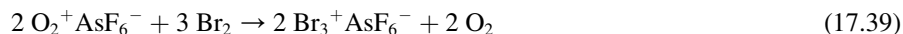
Although the cations I<sub>2</sub><sup>+</sup>, Br<sub>2</sub><sup>+</sup>, and Cl<sub>2</sub><sup>+</sup> are known, the triatomic species I<sub>3</sub><sup>+</sup>, Br<sub>3</sub><sup>+</sup>, and Cl<sub>3</sub><sup>+</sup> are better known chemically and they have been more fully characterized. For example, I<sub>3</sub><sup>+</sup> is generated in 100% H<sub>2</sub>SO<sub>4</sub> by the reaction



The Br<sub>3</sub><sup>+</sup> ion is generated by the following reaction that occurs in the superacid system HSO<sub>3</sub>F/SbF<sub>5</sub>/SO<sub>3</sub> (see Section 6.3).



Also, a compound containing the Br<sub>3</sub><sup>+</sup> cation is obtained by a reaction of bromine with dioxygenyl hexafluoroarsenate(V), O<sub>2</sub>AsF<sub>6</sub>. The reaction can be represented as follows:



The corresponding Cl<sub>3</sub><sup>+</sup> cation has also been obtained utilizing the reaction



The product is stable at dry ice-acetone temperatures, but it decomposes at room temperature. The reaction of I<sub>2</sub> with I<sub>3</sub><sup>+</sup> results in the formation of I<sub>5</sub><sup>+</sup>.

### 17.4.2 Interhalogen Cations

Although numerous species are possible that have the formula XYZ<sup>+</sup> (in which X, Y, and Z are different halogens), the most important interhalogen cations have the general formula XY<sub>2</sub><sup>+</sup>. Of these, not all of the possible species are known. The interhalogen cations that are best characterized are ClF<sub>2</sub><sup>+</sup>, BrF<sub>2</sub><sup>+</sup>, IF<sub>2</sub><sup>+</sup>, and ICl<sub>2</sub><sup>+</sup>. All of these ions have bent (C<sub>2v</sub>) structures as expected for species having a total of four electron pairs around the central atom.

A small degree of autoionization of the XX<sub>3</sub> interhalogens is indicated by their electrical conductivity. Some of these compounds have been extensively used as nonaqueous solvents in which their behavior indicates dissociation as shown in the following case for BrF<sub>3</sub>:



However, such reactions are hardly suitable for generating sufficient concentrations of the cations for study. Therefore, the cations described above are all produced by other means. In each case, the positive ions are generated by reaction of the trihalide with a strong Lewis acid that is capable of removing a fluoride ion. The following reactions illustrate the processes that are capable of producing the cations listed above.

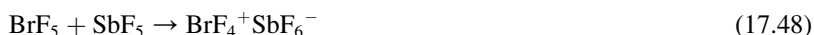


When  $\text{AsF}_5$  or  $\text{BF}_3$  reacts with  $\text{ClF}$ , adducts are obtained that appear to have the formulas  $2 \text{ClF} \cdot \text{AsF}_5$  or  $2 \text{ClF} \cdot \text{BF}_3$ . However, they can also be represented by the formulas  $\text{Cl}_2\text{F}^+\text{AsF}_6^-$  and  $\text{Cl}_2\text{F}^+\text{BF}_4^-$ , respectively. In these cases ionic species are present so they are more appropriately represented by the ionic formulas, and the cation  $\text{Cl}_2\text{F}^+$  is  $\text{ClClF}^+$  rather than  $\text{ClFCl}^+$ .

Crystal structures of compounds that contain the  $\text{ClF}_2^+$  ion show that the bond angle in the bent structure of this cation varies somewhat depending on the anion present. For example, in  $\text{ClF}_2^+\text{SbF}_6^-$  and  $\text{ClF}_2^+\text{AsF}_6^-$ , the  $\text{F}-\text{Cl}-\text{F}$  bond angles are  $95.9^\circ$  and  $103.2^\circ$ , respectively. This variation appears to be due to bridging between the chlorine atom in the  $\text{ClF}_2^+$  and fluorine atoms in the  $\text{SbF}_6^-$  and  $\text{AsF}_6^-$  anions. The bridged structure resulting from the chlorine atom interacting with two fluorine atoms in separate anions gives approximately a square plane of fluorine atoms around the chlorine atom in each case. Similar results have been found for other  $\text{XX}'_2^+$  ions. Because the conductivity of liquid  $\text{IF}_5$  is high enough to indicate some autoionization as represented by the equation



the cation  $\text{IF}_4^+$  is suggested. Similar species are produced from  $\text{ClF}_5$  and  $\text{BrF}_5$  by removal of  $\text{F}^-$  by a strong Lewis acid.



Iodine heptafluoride reacts with  $\text{AsF}_5$  or  $\text{SbF}_5$  to form products containing the  $\text{IF}_6^+$  cation.

### 17.4.3 Polyatomic Halogen Anions

It has been known for many years that solutions containing iodide ions dissolve  $\text{I}_2$  with the formation of  $\text{I}_3^-$ .



This process continues in the presence of excess  $\text{I}_2$  with the formation of higher anionic species that can be represented by the following equations:



The lighter halogens exhibit less tendency to form such polyatomic ions, although they do form some such species. For example, the reaction of chlorine with  $\text{Cl}^-$  forms the  $\text{Cl}_3^-$  ion,



Formation of these anionic species is not limited to species that contain only one halogen, and some interhalogen anions are also formed.



Ions such as  $\text{BrCl}_2^-$ ,  $\text{I}_3^-$ , and  $\text{Cl}_3^-$  are linear ( $D_{\infty h}$ ) whereas those such as  $\text{ICl}_4^-$  and  $\text{BrCl}_4^-$  are square planar ( $D_{4h}$ ) with two unshared pairs of electrons on the central atom. Stable compounds containing these polyhalide ions usually contain large cations such as  $\text{Rb}^+$ ,  $\text{Cs}^+$ , or tetraalkylammonium,  $\text{R}_4\text{N}^+$ , as predicted by the hard-soft interaction principle (see Chapter 6). As in the case of polyhalide cations, the less electronegative halogen is found in the central position in accord with it having a positive formal charge. The most thoroughly studied ions of this interhalogen type are  $\text{IBr}_2^-$ ,  $\text{ICl}_2^-$ , and  $\text{BrICl}^-$ .

The equilibrium constant for the reaction



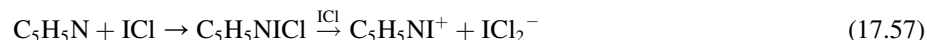
has been shown to be strongly dependent on the solvent used. For example in  $\text{H}_2\text{O}$ ,  $K = 16.3$ ; in a 50–50 weight percent mixture of  $\text{CH}_3\text{OH}$  and  $\text{H}_2\text{O}$ ,  $K = 58$ ; in 100%  $\text{CH}_3\text{OH}$ ,  $K = 176$ . This would be expected on the basis of the stronger solvation of  $\text{Br}^-$  by the more polar and smaller  $\text{H}_2\text{O}$  molecules that would result in a lower tendency for these ions to associate with  $\text{Br}_2$  molecules.

Solutions of iodine in organic solvents illustrate an important aspect of solution chemistry. Iodine vapor has a deep purple color as a result of absorption of light in the visible region which shows a maximum at 538 nm. Solutions of iodine

have absorption maxima that vary with the nature of the solvent, and when I<sub>2</sub> is dissolved in a solvent such as CCl<sub>4</sub> or heptane the solution has a blue-purple color. However, if the solvent is benzene or an alcohol the solution is brown. This difference in color is the result of the change in the relative energy of the  $\pi$  and  $\pi^*$  orbitals of the I<sub>2</sub> molecule as a result of their interaction with the solvent. The interaction is very weak for solvents such as CCl<sub>4</sub> and heptane so in those solvents the absorption maximum is close to where it is for gaseous I<sub>2</sub>. However, I<sub>2</sub> is a Lewis acid that interacts with electron pair donors. As a result, molecules that have unshared pairs of electrons such as alcohols or accessible electrons as in the case of the  $\pi$  system in benzene interact with I<sub>2</sub> molecules to perturb the molecular orbitals. The absorption is shifted to lower energy the more strongly the solvent interacts with the I<sub>2</sub> molecules. A change in the absorption spectrum (and hence a change in color) produced by a solvent is referred to as *solvatochromism*. A similar change in color that results from a change in temperature is referred to as *thermochromism*.

The nature and magnitude of the solvent–solute interaction depend on the molecular structures of the species. However, it should be apparent that this type of interaction provides a way to assess the interaction between a solute and the solvent. This is an extremely important area of chemistry with regard to understanding the role of the solvent as it relates to effects on solubility, equilibria, spectra, and rates of reactions. As a result, several numerical scales have been devised to correlate the effects of solvent interactions, some of which are based on solvatochromic effects. However, in most cases complex dyes have been utilized as the probe solutes, but it is interesting to note that iodine also exhibits solvatochromism.

The Lewis acid behavior of I<sub>2</sub> is well known (see Chapter 6). Interhalogens are also capable of behaving in this way, and ICl and IBr form complexes with bases such as pyridine.



Reactions such as this are strongly solvent dependent because part of the driving force must come from the solvation of the ionic species that are produced. For example, C<sub>5</sub>H<sub>5</sub>NICl does not ionize significantly in CHCl<sub>3</sub>. It has been suggested that *N*-iodo-pyridinium bromide undergoes ionization in a different way as shown in the following equation:



In aqueous HBr, the products, which are identified spectrophotometrically, have been found to be different depending on the concentration of acid. In very dilute acid, the product obtained from pyridine and IBr is C<sub>5</sub>H<sub>5</sub>NIBr. In acid concentrations above 1 M, the major product is C<sub>5</sub>H<sub>5</sub>NH<sup>+</sup>IBr<sub>2</sub><sup>−</sup>.

## 17.5 HYDROGEN HALIDES

### 17.5.1 Physical Properties

The physical properties of the hydrogen halides show the trends expected for increasing molecular weight except for HF which is anomalous owing to strong hydrogen bonding. In the condensed phases, liquid HF is extensively associated, and some aggregates exist in the vapor. As in the case of H<sub>2</sub>O, molecular association causes HF to have an extended liquid range that allows it to be useful as a nonaqueous solvent. Liquid HF is a good solvent for ionic compounds because it has a large dipole moment and a high dielectric constant. Some of the important properties of hydrogen halides are summarized in Table 17.3.

**TABLE 17.3** Properties of Hydrogen Halides

	HF	HCl	HBr	HI
m.p., °C	−83	−112	−88.5	−50.4
b.p., °C	19.5	−83.7	−67	−35.4
Bond length, pm	91.7	127.4	141.4	160.9
Dipole moment, Debye	1.74	1.07	0.788	0.382
Bond energy, kJ mol <sup>−1</sup>	574	428	362	295
$\Delta H_f^\circ$ , kJ mol <sup>−1</sup>	−273	−92.5	−36	+26
$\Delta G_f^\circ$ , kJ mol <sup>−1</sup>	−271	−95.4	−53.6	+1.6

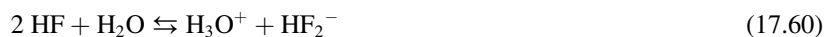
Liquid HF is miscible with H<sub>2</sub>O in all proportions. The other hydrogen halides are very soluble and form constant boiling mixtures with water. Because of this, constant boiling HCl has been used as a primary standard for analytical purposes. At 760 torr, the compositions and boiling points of the constant boiling mixtures are as follows:

Acid	Wt% HX	Boiling Point, °C
HCl	20.221	108.6
HBr	47.63	124.3
HI	53	127

Although the hydrogen halides all dissolve in water to give acidic solutions, there is a great difference in the acidity. Hydrogen fluoride is a weak acid whereas all the others are strong. The  $pK_a$  values are 2.92, -7, -9, and -9.5 for HF, HCl, HBr, and HI, respectively. The weakly acidic character of HF is due in part to the fact that F<sup>-</sup> is a hard base and competes effectively with H<sub>2</sub>O for the protons (see Chapter 6). Consequently, the reaction



does not proceed very far to the right. However, the reaction is not as simple as represented in this equation and other species are present in the solution. They are the result of reactions that can be shown as follows:



These types of processes are not, of course, of any consequence for the other hydrogen halides but rather are the result of the stability of the HF<sub>2</sub><sup>-</sup> ion. The most important chemical property of the hydrogen halides is, nevertheless, their acidity.

### 17.5.2 Preparation

The preparation of the hydrogen halides can be carried out by several methods. Not all of these methods will work for all of the compounds, however. Because of their volatility, the usual method for preparing HF, HCl, or HBr is by heating a halide salt with a nonvolatile acid such as H<sub>2</sub>SO<sub>4</sub> or H<sub>3</sub>PO<sub>4</sub>.

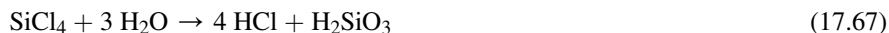


As a result of the ease with which the iodide compound dissociates at elevated temperature, this method is not applicable to the preparation of HI. Also, if an acid is used that can behave as an oxidizing agent, the reducing power of HI leads to a redox reaction producing the halogen.



Sulfuric acid is also reduced by HBr if the H<sub>2</sub>SO<sub>4</sub> is both hot and concentrated.

It is possible to prepare the hydrogen halides by direct combination of the elements although this method is difficult to control (especially in the case of HF and HCl) or does not give rapid, efficient reaction (as with HI). A third method, the preferred one for HI, is the hydrolysis of covalent halogen compounds containing iodine bonded to a nonmetal.



In the case of HCl, large quantities are produced industrially by the chlorination of hydrocarbons.

## 17.6 OXIDES

Several oxides of the Group VIIA elements are known, but only those of chlorine are of any commercial importance. Most are unstable, some explosively so. Fluorine is more electronegative than oxygen so the compounds containing oxygen and

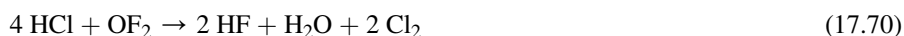
fluorine are correctly considered to be fluorides. Because oxygen has a higher electronegativity than Cl, Br, and I, the oxygen compounds of these elements are oxides.

### 17.6.1 Oxygen Fluorides

The simplest and best-known compound of this type is oxygen difluoride, OF<sub>2</sub> (b.p. 145 °C). It is a pale yellow, poisonous gas that can be prepared by passing fluorine through a dilute solution of NaOH or by electrolyzing aqueous solutions containing KF and HF. It has a bent (C<sub>2v</sub>) structure with a bond angle of 103.2° and an O–F bond length of 141 pm. The compound reacts with water as shown in the following equation:



Reactions of OF<sub>2</sub> with hydrogen halides or halide salts produce the free halogen. For example,



Above 250 °C, OF<sub>2</sub> decomposes into O<sub>2</sub> and F<sub>2</sub>.

Dioxygen difluoride, O<sub>2</sub>F<sub>2</sub>, is a yellow-orange solid melting at –163 °C. It is prepared by the reaction of O<sub>2</sub> and F<sub>2</sub> mixtures under glow discharge at –180 to –190 °C. The structure of the molecule is similar to that of H<sub>2</sub>O<sub>2</sub> (see Chapter 15), and it has an O–O distance of 122 pm, an O–F distance of 157 pm, and a dihedral angle of 87.5°. The compound is an extremely active fluorinating agent. Although O<sub>3</sub>F<sub>2</sub> and O<sub>4</sub>F<sub>2</sub> are also produced during glow discharge in mixtures of O<sub>2</sub> and F<sub>2</sub>, they are very unstable and decompose even at liquid nitrogen temperatures.

### 17.6.2 Chlorine Oxides

Several chlorine oxides are known, and they are the anhydrides of the chlorine oxyacids. One of them, ClO<sub>2</sub>, has several important uses on an industrial scale. Dichlorine monoxide, Cl<sub>2</sub>O is an explosive gas (m.p. –116 °C, b.p. +2 °C) that is obtained as a product of the following reaction:



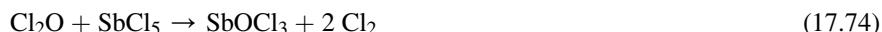
This oxide is very soluble in water, and it is the anhydride of HOCl.



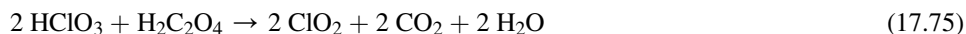
This reaction illustrates the major use of Cl<sub>2</sub>O. The Cl<sub>2</sub>O molecule has a bent (C<sub>2v</sub>) structure with a Cl–O–Cl angle of 110.8°. It reacts with N<sub>2</sub>O<sub>5</sub> to give chlorine nitrate, ClNO<sub>3</sub>.



The other products in Eq. (17.73) depend on the reaction conditions. In the ClNO<sub>3</sub> molecule the chlorine atom is bonded to oxygen so the linkage is ClONO<sub>2</sub>. Dichlorine monoxide also reacts with metal chlorides, and the reaction with SbCl<sub>5</sub> can be represented by the equation



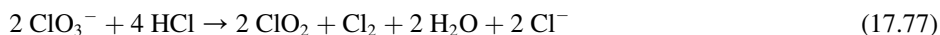
Chlorine dioxide, ClO<sub>2</sub>, is an explosive gas (m.p. –60 °C, b.p. 11 °C) that has been used as a dry, gaseous bleach. It has a bent (C<sub>2v</sub>) structure with an O–Cl–O angle of 118°. Although ClO<sub>2</sub> has an odd number of electrons, it does not dimerize. The reaction,



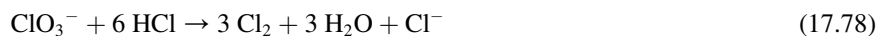
can be used to prepare ClO<sub>2</sub>. This oxide reacts in basic solutions to produce a mixture of ClO<sub>2</sub><sup>–</sup> and ClO<sub>3</sub><sup>–</sup>.



Industrially, ClO<sub>2</sub> is prepared by several methods, most of which involve reduction of NaClO<sub>3</sub>. For example, the reaction with HCl is



Because of competing reactions one of which is represented by the following equation,



under certain conditions the reaction is different, and the overall process approaches that shown in Eq. (16.79).



This situation illustrates why there are cases where it is often difficult to predict the products of a reaction with certainty. Catalysts for the reduction include  $\text{Ag}^+$ ,  $\text{Mn}^{2+}$ ,  $\text{V}_2\text{O}_5$ , and peroxides.

Reduction of  $\text{ClO}_3^-$  using  $\text{SO}_2$  is also used to prepare  $\text{ClO}_2$ .



Another commercial method of preparing  $\text{ClO}_2$  involves the reduction of  $\text{ClO}_3^-$  with gaseous methanol as shown by the following equation:

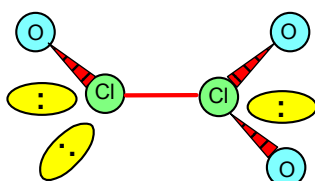


Also,  $\text{ClO}_2$  results from the reaction of solid  $\text{NaClO}_3$  with gaseous chlorine.



There are several important industrial uses of  $\text{ClO}_2$ . Most of these involve its bleaching action, but it is also a useful bactericide in water treatment. Bleaching wood pulp, textiles, and peat are some of the industrial processes in which  $\text{ClO}_2$  is employed. Large quantities of  $\text{ClO}_2$  are used as a gaseous, dry bleach for flour. For this and other uses, it is usually generated at the place of use. Although it is used extensively (with great caution),  $\text{ClO}_2$  is a dangerous explosive having a heat of formation of  $+105 \text{ kJ mol}^{-1}$ .

When  $\text{ClO}_2$  is photolyzed at  $-78^\circ\text{C}$ , it apparently yields  $\text{Cl}_2\text{O}_3$  as one product. The structure of the  $\text{Cl}_2\text{O}_3$  molecule is



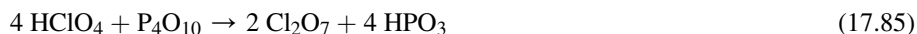
An unusual compound known as chlorine perchlorate,  $\text{ClOClO}_3$ , can be produced by the following reaction:



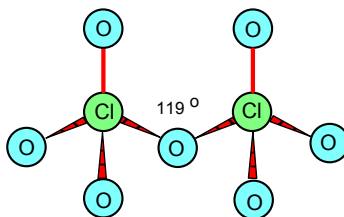
However, this compound is not a useful oxide of chlorine. The reaction of ozone with  $\text{ClO}_2$  can be shown as follows:



The product, dichlorine hexoxide, behaves as  $[\text{ClO}_2^+][\text{ClO}_4^-]$  in some of its reactions and in the solid state. Dehydration of  $\text{HClO}_4$  using  $\text{P}_4\text{O}_{10}$  produces  $\text{Cl}_2\text{O}_7$  that can be separated by distillation at low pressures and temperatures.



The molecular structure of  $\text{Cl}_2\text{O}_7$  is

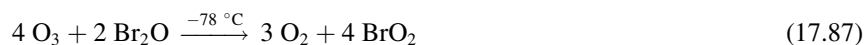


### 17.6.3 Bromine Oxides

Although some bromine oxides have been studied, all appear to be unstable at temperatures above  $-40^\circ\text{C}$ . The monoxide is prepared by the reaction of bromine with mercuric oxide,



Also, as shown by the following equation, the reaction of  $\text{Br}_2\text{O}$  with ozone produces  $\text{BrO}_2$ .



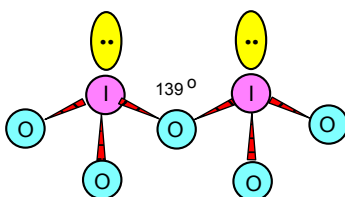
Unlike  $\text{ClO}_2$ , neither of the bromine oxides described above is of much importance.

### 17.6.4 Iodine Oxides

The most extensively studied oxide of iodine is  $\text{I}_2\text{O}_5$ . It is prepared by the dehydration of iodic acid,



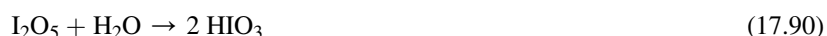
and its structure is



The white solid decomposes at high temperatures ( $>300^\circ\text{C}$ ) to give  $\text{I}_2$  and  $\text{O}_2$ . Chemically,  $\text{I}_2\text{O}_5$  is a strong oxidizing agent and this property has been extensively exploited. In one use, it is employed to oxidize  $\text{CO}$  for quantitative determination.



This oxide reacts with water and it is the anhydride of iodic acid,



The only other oxides of iodine that appear to have been prepared are  $\text{I}_2\text{O}_4$ ,  $\text{I}_4\text{O}_9$ , and  $\text{I}_2\text{O}_7$ , but they are of no real importance.

### 17.6.5 Oxyfluorides of the Heavier Halogens

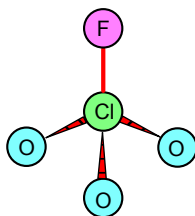
In Chapter 16, some of the chemistry of the oxyhalides of S, Se, and Te was described. The Group VIIA elements (except F) form similar compounds in which a halogen is the central atom. Most of these compounds are fluorides, as are most of the interhalogens, and they are strong oxidizing and/or fluorinating agents. Known compounds include  $\text{FCIO}_2$ ,  $\text{F}_3\text{ClO}$ ,  $\text{FCIO}_3$ ,  $\text{F}_3\text{ClO}_2$ ,  $\text{ClO}_3\text{OF}$ ,  $\text{FBrO}_2$ ,  $\text{FBrO}_3$ ,  $\text{FIO}_2$ ,  $\text{F}_3\text{IO}$ ,  $\text{FIO}_3$ ,  $\text{F}_3\text{IO}_2$ , and  $\text{F}_5\text{IO}$ . The hydrolysis reactions of these compounds produce the corresponding acids in most cases. An example of this type of behavior is illustrated for  $\text{FCIO}_2$ .



The compounds also react with Lewis acids to generate cationic species by fluoride ion donation.



Of these compounds,  $\text{FCIO}_3$ , perchloryl fluoride, is the most stable. It is stable to  $\sim 500^\circ\text{C}$  and it does not hydrolyze rapidly. The  $\text{FCIO}_3$  molecule has a  $C_{3v}$  structure.



Because of its oxidizing strength, it has been used as the oxidant with suitable rocket fuels.

## 17.7 OXYACIDS AND OXYANIONS

A relatively large number of oxyacids of the halogens exist. Because of the extensive redox chemistry of these compounds, there is an enormous number of reactions involving them. We will give only an overview in keeping with the spirit and space requirements of this survey.

### 17.7.1 Hypohalous Acids and Hypohalites

All of the halogens form acids having the formula HOX, and all are oxidizing agents. Of these, only HOF appears to be stable outside of aqueous solutions, and it decomposes easily into HF and O<sub>2</sub>. The HOX acids can all be prepared by the general reaction



that is driven to the right by the removal of X<sup>-</sup> as a precipitate with an appropriate cation. The HOX acids are weak having K<sub>a</sub> values of 3 × 10<sup>-8</sup>, 2 × 10<sup>-9</sup>, and 1 × 10<sup>-11</sup> for HOCl, HOBr, and HOI, respectively.

These acids can also be prepared by the reaction of the appropriate halogen with a base,



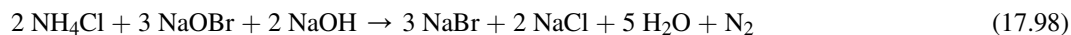
However, the solutions must be cold and dilute or else the competing reaction



takes place. Therefore, the actual reactions between halogens and basic solutions are dependent on temperature, pH, and concentrations. The acids HOX are unstable and decompose in reactions that can be represented as follows:



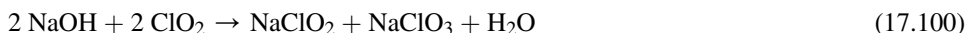
The decomposition of OCl<sup>-</sup> is catalyzed by transition metal compounds. For example, adding aqueous Co(NO<sub>3</sub>)<sub>2</sub> to a solution of NaOCl results in very rapid decomposition of the hypochlorite with the liberation of oxygen and produces a solid catalyst that is probably CoO and/or Co<sub>2</sub>O<sub>3</sub>. Hypochlorites are extensively used as bleaching agents in many processes. They are also versatile oxidizing agents that oxidize CN<sup>-</sup> to OCN<sup>-</sup>, AsO<sub>3</sub><sup>3-</sup> to AsO<sub>4</sub><sup>3-</sup>, NO<sub>2</sub><sup>-</sup> to NO<sub>3</sub><sup>-</sup>, etc. Hypobromites and hypoiodites are also good oxidizing agents but are less extensively used. NaOBr has been used in analytical procedures to oxidize NH<sub>4</sub><sup>+</sup> and urea, (NH<sub>2</sub>)<sub>2</sub>CO, as shown in the following equations:



### 17.7.2 Halous Acids and Halites

In the case of fluorine, the only oxyacid known is HOF. For acids having the formula HXO<sub>2</sub> (or more accurately, HOXO), only HClO<sub>2</sub> is well characterized and even for that acid only aqueous solutions are known. Consequently, the discussion will be limited to chlorous acid and its salts.

Sodium chlorite can be prepared by the reaction of ClO<sub>2</sub> with aqueous NaOH.

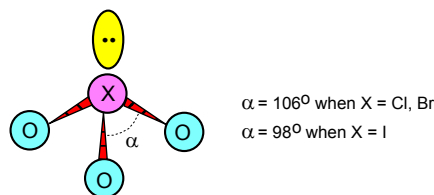


The chlorite ion is angular (C<sub>2v</sub>) with a bond angle of about 111°. A solution of HClO<sub>2</sub> can be prepared by acidifying a chlorite solution with H<sub>2</sub>SO<sub>4</sub>. In accord with b = 1 in the general formula (HO)<sub>a</sub>XO<sub>b</sub>, the acids are weak (see Chapter 6).

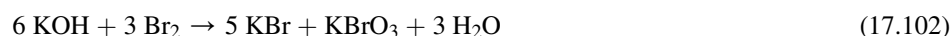
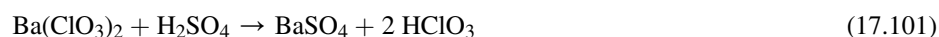


### 17.7.3 Halic Acids and Halates

These acids, having the formula  $(\text{HO})_a\text{XO}_b$  with  $b = 2$ , are strong acids, but  $\text{HClO}_3$  and  $\text{HBrO}_3$  exist only as aqueous solutions. The  $\text{XO}_3^-$  anions have 26 electrons, leading to the  $C_{3v}$  structure



The acids are obtained by several reactions including the following:



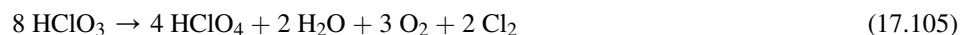
For the preparation of  $\text{BrO}_3^-$ , the following reaction can be employed:



Commercially, electrolysis of a solution of  $\text{NaCl}$  is used to prepare  $\text{NaClO}_3$ .



The acids  $\text{HClO}_3$  and  $\text{HBrO}_3$  decompose when heated, and the disproportionation reaction of  $\text{HClO}_3$  produces the perchlorate.



However, in the case of  $\text{HBrO}_3$  decomposition occurs as shown in the following equation:



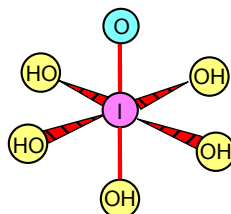
Chlorates, bromates, and iodates are all strong oxidizing agents that oxidize many materials. Iodic acid can be dehydrated to produce  $\text{I}_2\text{O}_5$ .

### 17.7.4 Perhalic Acids and Perhalates

Perchloric and perbromic acids are very strong, having  $b = 3$  in the general formula  $(\text{HO})_a\text{XO}_b$ . Periodic acid in an excess of water is a weak acid because the  $\text{HIO}_4$  initially present reacts with water to give  $\text{H}_5\text{IO}_6$ .



The  $\text{H}_5\text{IO}_6$  molecule has a structure ( $C_{4v}$ ) that can be shown as



thereby giving  $b = 1$  in the general formula  $(\text{HO})_a\text{XO}_b$ . This behavior is similar to that shown by  $\text{Te}$  (see Chapter 16). A series of acids can be postulated as resulting from the reaction of  $\text{I}_2\text{O}_7$  with  $\text{H}_2\text{O}$ .



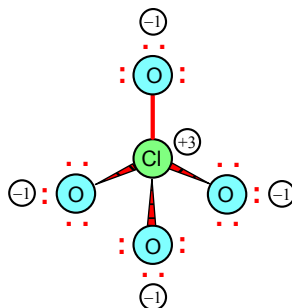


Thus, in aqueous solutions a very complicated series of acidic species results.

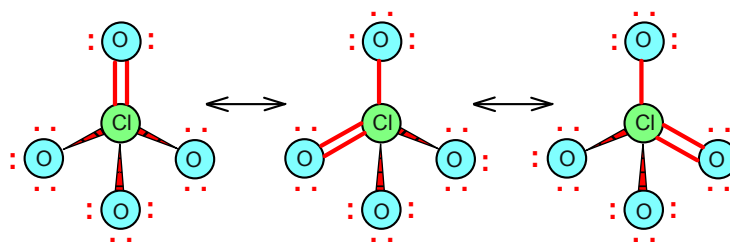
The acid  $\text{HIO}_4$  is known as *periodic* or *metaperiodic* acid whereas  $\text{H}_5\text{IO}_6$  is known as *orthoperiodic* acid or sometimes *paraperiodic* acid. The prefixes have nothing to do with their usual meaning in organic chemistry. Note the similarity to the phosphoric acid names (see Chapter 14).

Perbromic acid is stable only in aqueous solutions, although  $\text{HClO}_4$  can be obtained by distillation under reduced pressure using a suitable drying agent. The pure  $\text{HClO}_4$  may decompose explosively, and contact with oxidizable organic materials may cause explosions. The usual commercial form of  $\text{HClO}_4$  is a solution that contains 70% of the acid by weight. Solid metal perchlorates may also explode in contact with organic materials. Perchloric acid is obtained by treating solutions of metal perchlorates with a strong acid and removing the metal salt by filtration. The metal perchlorates are usually obtained by electrolytic oxidation of a chloride in aqueous solution or by heating a metal chlorate that results in disproportionation. Perbromic acid was first obtained in 1968 by the  $\beta^+$ -decay of  $^{83}\text{Se}$  in  $^{83}\text{SeO}_4^{2-}$ . Also, electrolytic oxidation of  $\text{BrO}_3^-$  or oxidation by  $\text{F}_2$  or  $\text{XeF}_2$  can be used. Periodates are prepared by oxidation of  $\text{I}^-$  or  $\text{IO}_3^-$ . The periodic acid is prepared by the reaction of  $\text{Ba}_5(\text{IO}_6)_2$  with concentrated  $\text{HNO}_3$  that results in the formation of  $\text{H}_5\text{IO}_6$ .

If one considers the tetrahedral ( $T_d$ ) structure of  $\text{ClO}_4^-$ ,



it is apparent that the formal charge on the chlorine is +3. In order to alleviate this situation, some multiple bonding occurs as a result of the donation of electron density from the filled  $p$  orbitals of the oxygen atoms to the empty  $3d$  orbitals on the chlorine atom. A similar situation was described for the  $\text{SO}_4^{2-}$  structure (see Chapter 16). The observed bond distance in  $\text{ClO}_4^-$  is somewhat shorter than expected for a single  $\text{Cl}-\text{O}$  bond because structures such as



make substantial contributions to the actual structure.

This chapter has surveyed the enormously varied chemistry of the halogens. However, this is such a broad topic that entire volumes have been devoted to the chemistry of this fascinating group of elements. The reader interested in going beyond the coverage presented in this chapter should consult the references listed.

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## PROBLEMS

- Explain what differences would exist between the properties of a solution made by dissolving  $\text{Cl}_2\text{O}_7$  in water and dissolving  $\text{I}_2\text{O}_7$  in water. What is the origin of the differences?
- Draw structures for each of the following species.
  - $\text{BrF}_4^+$
  - $\text{Cl}_2\text{O}_7$
  - $\text{IBr}_2^-$
  - $\text{IF}_5$
  - Periodic acid
  - $\text{IOF}_5$
- Write balanced equations for each of the following processes.
  - Reaction of iodine trichloride with water
  - Preparation of  $\text{NaOCl}$
  - Preparation of  $\text{KI}_3$
  - Disproportionation of  $\text{OCl}^-$
  - Electrolysis of a dilute sodium chloride solution
- On the basis of energy, explain why fluorine is an even more reactive element than might be expected when compared to the reactivity of the other halogens.
- Iodine dissolved in carbon tetrachloride is purple. In acetone, the color is brown. Explain these observations.
- Discuss the reasons why  $\text{HF}$  is a weak acid.
- Draw the molecular orbital energy level diagram for  $\text{NO}$ . What indication do you see for the observed chemistry of the molecule as it pertains to halogen chemistry?
- Write an equation to show how  $\text{SbF}_5$  behaves as an acid in liquid  $\text{BrF}_3$ .
- Complete and balance the following.
  - $\text{MnO}_2 + \text{HCl} \xrightarrow{\text{heat}}$
  - $\text{ClF} + \text{H}_2\text{O} \rightarrow$
  - $\text{BrF}_3 + \text{AsF}_5 \rightarrow$
  - $\text{H}_2\text{O} + \text{OF}_2 \rightarrow$
  - $\text{HClO}_3 + \text{P}_4\text{O}_{10} \rightarrow$
- Give formulas for three species that are isoelectronic with  $\text{Cl}_2\text{O}_7$  that do not contain halogen atoms.
- Complete and balance the following.
  - $\text{BrF}_3 + \text{H}_2\text{O} \rightarrow$
  - $\text{S} + \text{ClF}_3 \rightarrow$
  - $\text{HI} + \text{H}_2\text{SO}_4 \rightarrow$
  - $\text{ONF} + \text{ClF} \rightarrow$
- Draw structures for the following species.
  - $\text{ClF}_2^+$
  - $\text{I}_3^-$
  - $\text{ClO}_2^-$
  - $\text{I}_2\text{O}_7$
  - $\text{Br}_3^+$

13. Complete and balance the following.

- (a)  $\text{Ca}(\text{OH})_2 + \text{Cl}_2 \rightarrow$
- (b)  $\text{KMnO}_4 + \text{HCl} \rightarrow$
- (c)  $\text{Cl}_2\text{O} + \text{H}_2\text{O} \rightarrow$
- (d)  $\text{NaOH} + \text{I}_2 \rightarrow$
- (e)  $\text{SO}_2\text{Cl}_2 + \text{H}_2\text{O} \rightarrow$

14. Complete and balance the following.

- (a)  $\text{F}_2 + \text{H}_2\text{O} \rightarrow$
- (b)  $\text{ICl} + \text{H}_2\text{O} \rightarrow$
- (c)  $\text{KF} + \text{IF}_5 \rightarrow$
- (d)  $\text{SiO}_2 + \text{ClF}_3 \rightarrow$
- (e)  $\text{PI}_3 + \text{H}_2\text{O} \rightarrow$

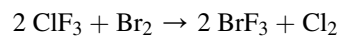
15. Draw structures for the following.

- (a)  $\text{IF}_6^-$
- (b)  $\text{ClF}_2^-$
- (c)  $\text{IF}_4^+$
- (d)  $\text{OF}_2$
- (e)  $\text{BrF}_4^-$

16. Explain in terms of periodic properties why compounds such as  $\text{IClO}_4$  and  $\text{I}(\text{py})_2\text{ClO}_4$  (py = pyridine) are much more stable than those with Cl replacing I.

17. For a cation containing two chlorine atoms and one fluorine atom ( $\text{Cl}_2\text{F}^+$ ), two structures are possible. Explain the difference in stability expected for these ions.

18. Although chlorine is a stronger oxidizing agent than bromine, the following reaction takes place.



Explain this observation.

19. Although  $\text{ClO}_4^-$  does not coordinate to metal ions strongly, numerous complexes are known in which it is a ligand. Describe ways in which  $\text{ClO}_4^-$  could form coordinate bonds to metal ions.

20. The I—O (the oxygen atom that does not have a hydrogen atom attached) bond length in  $\text{H}_5\text{IO}_6$  is 178 pm. Assume that a proton is removed to form the  $\text{H}_4\text{IO}_6^-$  ion. Which proton would likely be removed? In this ion, the length of the bonds between I and the *two* oxygen atoms with no hydrogen atom attached is 186 pm. Explain why this is different from the value of 178 pm in  $\text{H}_5\text{IO}_6$ .

# The Noble Gases

For many years, the name “inert gases” was applied to the elements in Group VIIIA, but that name has been shown to be incorrect. Consequently, the gases are now known as the *noble* gases. Because of their closed shell electron configurations and high ionization potentials, it was long believed that they were incapable of forming chemical compounds. Cage type structures were known in which the gaseous atoms were encapsulated, but such materials tend to be nonstoichiometric and without chemical bonds in the usual sense. For example, freezing water that is under pressure by xenon causes some of the atoms to be trapped in ice “cages” that can be represented by the formula  $\text{Xe}(\text{H}_2\text{O})_n$ . Molecules trapped inside other compounds (and there are many) are called *clathrates*.

All of this changed when the reaction

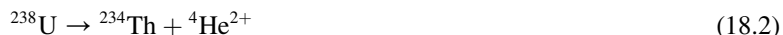


was carried out by Neil Bartlett and D. H. Lohmann in 1962. In this case, the extreme oxidizing power (ability to remove electrons) of  $\text{PtF}_6$  was exploited. The ionization potential for the  $\text{O}_2$  molecule is approximately  $1177 \text{ kJ mol}^{-1}$  whereas that for the xenon atom is about  $1170 \text{ kJ mol}^{-1}$ . These observations suggested that perhaps xenon would react with  $\text{PtF}_6$  to form a compound similar to that produced by oxygen. Such experiments were carried out, and an orange-yellow solid was obtained. The initial account indicated that the solid, which is insoluble in carbon tetrachloride, was  $\text{Xe}^+\text{PtF}_6^-$ , although later work indicated that the initial compound was probably  $\text{XeF}^+[\text{Pt}_2\text{F}_{11}]^-$ . In this manner, the *chemistry* of the noble gases began in 1962. Later that year, the xenon fluorides  $\text{XeF}_4$ ,  $\text{XeF}_2$ , and  $\text{XeF}_6$  were reported. There are as yet no compounds of helium or neon, and only one compound containing argon,  $\text{HArF}$ , but compounds containing krypton and radon are known in addition to those of xenon. However, even though there is some chemistry of krypton and radon, there is a much more extensive chemistry of xenon as the remainder of this chapter will indicate.

## 18.1 THE ELEMENTS

Although the chemistry of the noble gases has been of relatively recent development, the noble gases themselves have been known for many years. For example, radon, which occurs only as a product of radioactive decay, has been known for over a century.

Helium was identified in the spectrum of the sun on the basis of its spectral lines by Lockyer and Frankland in 1868. In 1894, helium was found in uranium ore where it is produced by the decay of  $^{238}\text{U}$  as shown by the following equation:



The  $\alpha$  particles (helium nuclei) acquire two electrons to become helium atoms, some of which remain trapped in the ores. The Greek word *helios* for sun is the origin of the name helium.

Argon had already been discovered in the residue that remains after oxygen and nitrogen are removed from air. It was identified in 1885 by Sir William Ramsay on the basis of the lines emitted when the atoms are excited by electric discharge. It was isolated in 1894 by Lord Rayleigh and Sir William Ramsay with the name coming from the Greek word *argos* meaning inactive. Some naturally occurring argon is produced in the decay of  $^{40}\text{K}$  by electron capture (represented as E.C. in the following equation).



Neon, krypton, and xenon were isolated in 1898 from liquid air by Sir William Ramsay and M. W. Travers. The name for xenon comes from the Greek *xenos* meaning “strange.” The Greek word *kryptos* meaning “hidden” is the source of the name krypton and the Greek word *neos* meaning “new” is the source of the name for neon.

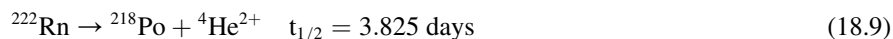
The gaseous element radon was first detected in 1899 by Ernest Rutherford (while studying thorium) and Pierre and Marie Curie (while studying radium). Friedrich Dorn is credited with discovering radon in 1900 when he detected a

radioactive gas that was accumulating in tubes of radium. Radon was originally named *nitron*, which comes from the Latin word *nitrens* that means “shining.”

Radium is formed when nuclei of certain naturally occurring heavy elements undergo decay by a series of steps that eventually leads to a stable product. The three radioactive decay series that occur naturally are those of  $^{235}\text{U}$ ,  $^{238}\text{U}$ , and  $^{232}\text{Th}$ , and radon is produced in each of these schemes as decay leads from the initial member of the series to a stable nuclide. However, all of the isotopes of radon are radioactive so they undergo further decay, which limits the chemistry of this element. In a sequence of steps, the decay of  $^{232}\text{Th}$  eventually leads to  $^{224}\text{Ra}$ , and it undergoes  $\alpha$ -decay to produce  $^{220}\text{Rn}$  for which the half-life is 54.5 s. As  $^{235}\text{U}$  decays in a scheme of many steps to produce  $^{207}\text{Pb}$  (a stable product),  $^{223}\text{Ra}$  is an intermediate product. It decays by  $\alpha$  emission to produce  $^{219}\text{Rn}$  that undergoes decay by  $\alpha$  emission with a half-life of only 3.92 s. During the several steps in the decay of  $^{238}\text{U}$  that leads to the final product  $^{206}\text{Pb}$ ,  $^{226}\text{Ra}$  is produced, and it decays by  $\alpha$  emission to produce  $^{222}\text{Rn}$  that has a half-life of 3.825 days. Although the entire decay schemes will not be shown, the relevant steps that lead to the production of radon can be summarized as follows:



The radioactive decay of the three radon isotopes after they are produced can be shown as follows:



Moreover, all of the products of these decay reactions are also radioactive.

The isotopes having very short half-lives do not normally present much of a health problem, but  $^{222}\text{Rn}$  is of great concern. Because of its very high atomic mass, radon is a very dense gas ( $9.73 \text{ g l}^{-1}$  compared to the density of  $1.29 \text{ g l}^{-1}$  for air at  $0^\circ\text{C}$  and 1 atm) so it tends to accumulate in low places such as basements and mines. Moreover, because of its high atomic mass, radon does not diffuse rapidly so it does not escape from these areas as would a lighter gas. The short half-life of two of the radon isotopes results in their not lasting very long, but  $^{222}\text{Rn}$  presents a problem. Ordinarily,  $\alpha$  and  $\beta$  particles are not much of a health hazard because they are not very penetrating and are easily stopped by even a layer of skin. However, when  $^{222}\text{Rn}$  undergoes radioactive decay after being inhaled, the  $^{218}\text{Po}$  and subsequent decay products are solids that are deposited on the lining of the lungs. When nuclei that have resulted from the decay of radon subsequently undergo radioactive decay, they do so in contact with sensitive lung tissue that incurs radiation damage even though  $\alpha$  and  $\beta$  radiation are not highly penetrating. One of the results of exposure to radiation in this way can be development of lung cancer, and the implication of radon as a cause is the subject of a great deal of research as well as press coverage. As usual, it is also an opportunity for a great deal of entrepreneurial effort by the makers of home testing kits for detecting radon and those who make a business of conducting tests for radon.

Because the noble gases represent elements that are of limited chemical reactivity, there is a progression in physical properties that reflects the increasing atomic weight of the gases. Several of the more important and relevant properties of the gases are shown in [Table 18.1](#).

One of the interesting characteristics of the noble gases is the very small temperature range of the liquid phase. Most of the noble gases have a liquid range that spans only four or five degrees in contrast to that of water for which the liquid range spans  $100^\circ\text{C}$ . A small liquid range is indicative of very weak forces of attraction between molecules and a random arrangement of molecules in the liquid. These characteristics are in accord with those expected for liquids composed of nonpolar molecules having only London dispersion forces of attraction. There being no other types of molecular interactions, the noble gases obey the ideal gas equation closely and are frequently cited as ideal gases.

Helium is used as a coolant for superconductors and, because of its low density, as a lifting gas for lighter than air aircraft. Neon is widely used in neon signs because it emits light when excited by electrical discharge. Neon is surprisingly abundant on a cosmic scale with an abundance that is greater than that of silicon. This is due in part to the fact that the nucleus of the most abundant isotope ( $^{20}\text{Ne}$ ) contains 10 protons and 10 neutrons so the two lowest nuclear energy levels are filled with protons and neutrons. A closed shell of neutrons and protons leads to nuclear stability in much the same way that a closed shell of electrons leads to chemical stability. Argon makes up approximately 1% of the atmosphere and is used

**TABLE 18.1** Some Properties of the Noble Gases

Property	He	Ne	Ar	Kr	Xe	Rn
m.p., K	0.95	24.5	83.78	116.6	161.3	202
b.p., K	4.22	27.1	87.29	120.8	166.1	211
$\Delta H_{\text{fusion}}$ , kJ mol <sup>-1</sup>	0.021	0.324	1.21	1.64	3.10	2.7
$\Delta H_{\text{vap}}$ , kJ mol <sup>-1</sup>	0.082	1.74	6.53	9.70	12.7	18.1
Ionization Energy, kJ mol <sup>-1</sup>	2372	2081	1520	1351	1170	1037
<sup>a</sup> Atomic radius, pm	122	160	191	198	218	~220
<sup>b</sup> Density, g L <sup>-1</sup>	0.1785	0.900	1.784	3.73	5.88	9.73

<sup>a</sup>van der Waals radii.<sup>b</sup>At 0°C and 1 atm.

to provide a nonoxidizing atmosphere during welding operations. In view of these uses and the fact that liquid air is used on an enormous scale to obtain liquid oxygen and nitrogen, the production of these noble gases as by-products is substantial.

## 18.2 THE XENON FLUORIDES

Not only were the fluorides the first compounds of xenon to be prepared, but also they serve as starting materials for the synthesis of most other xenon compounds. Xenon difluoride can be prepared by the reaction of excess xenon with fluorine aided by heat or electromagnetic radiation.



The preparation of xenon tetrafluoride is carried out by heating a 1:5 mixture of xenon to fluorine under a pressure of several atmospheres.

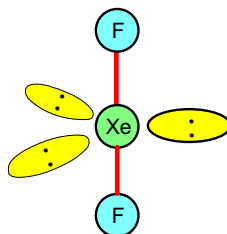


These conditions are also employed in the preparation of XeF<sub>6</sub> except that a larger excess of F<sub>2</sub> is used.

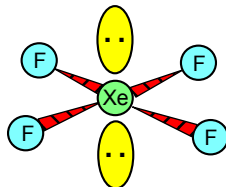
From a comparison of their ionization potentials, one would expect that krypton would form compounds with more difficulty than would xenon and this is the case. The difluoride of krypton is prepared by the reaction of krypton with fluorine under electric discharge at low temperature. It has also been prepared by the action of ultraviolet light on a mixture of liquid krypton and fluorine. The strength of the Xe–F bond is 133 kJ mol<sup>-1</sup> whereas that of the Kr–F bond is only 50 kJ mol<sup>-1</sup>. Therefore, KrF<sub>2</sub> is much less stable and more reactive than XeF<sub>2</sub>. Radon difluoride can also be prepared but because of the very short half-life of radon, much less experimental work has been carried out using the compound. Although compounds of krypton and radon are known, they are much less numerous and well studied than are those of xenon. Consequently, most of what is presented in this section applies to the xenon compounds.

Because of the high ionization potential of xenon, it is reasonable to assume that most of its compounds would contain bonds to atoms of high electronegativity such as F and O. Although this was initially so, there are now compounds known that contain Xe–Cl and Xe–N bonds.

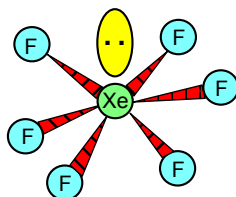
The structures of the binary fluorides are predictable on the basis of the valence shell electron pair repulsion model (see Chapter 3). With eight valence shell electrons from the xenon atom and two additional electrons from the two fluorine atoms, there are 10 electrons surrounding the xenon atom in XeF<sub>2</sub>. Thus, the structure of XeF<sub>2</sub> has *D*<sub>∞h</sub> symmetry as shown below.



The structure of  $\text{XeF}_4$  is derived from the fact that there are 12 electrons surrounding the Xe atom in the molecule (8 from the Xe atom and 1 from each fluorine atom). Therefore, the 12 electrons reside in 6 orbitals pointing toward the corners of an octahedron and the molecular structure is square planar ( $D_{4h}$ ).



The structure of  $\text{XeF}_6$  is not a regular octahedron owing to the fact that the xenon atom has 14 electrons around it. The  $\text{IF}_7$  molecule is similar in this regard and it has a pentagonal bipyramid structure ( $D_{5h}$ ). However, with one unshared pair of electrons, there is some question as to where the unshared pair resides. Also, the molecule is not rigid and it is described as an irregular octahedron having  $C_{3v}$  symmetry.



In the condensed phase, there is an equilibrium that can be shown as



in which there are fluoride bridges. There are also hexamers with fluoride bridges in the solid. Thus, the structure of  $\text{XeF}_6$  is not nearly as simple as that of  $\text{XeF}_2$  and  $\text{XeF}_4$ .

As was described in Chapter 4, intermolecular forces are related in a general way to the structures of the molecules and increase with molecular polarity. For nonpolar molecules, intermolecular forces tend to increase (leading to higher melting and boiling points) as the molecular weight of the molecules increase. Therefore, the fact that the melting points of  $\text{XeF}_2$ ,  $\text{XeF}_4$ , and  $\text{XeF}_6$  are 129, 117, and 50 °C, respectively, seems somewhat surprising.

In Chapter 4, it was shown that the solubility parameter can often provide some insight into the strength and nature of intermolecular forces. The solubility parameters calculated from vapor pressures of  $\text{XeF}_2$  and  $\text{XeF}_4$  are  $33.3 \text{ J}^{1/2} \text{ cm}^{-3/2}$  and  $30.9 \text{ J}^{1/2} \text{ cm}^{-3/2}$ , respectively. Both molecules are nonpolar, but if the molecules were interacting only by London dispersion forces we would expect a higher solubility parameter for  $\text{XeF}_4$  because of its higher molecular mass. For example, the solubility parameter for  $\text{SnCl}_4$  ( $260.5 \text{ g mol}^{-1}$ ) is  $17.8 \text{ J}^{1/2} \text{ cm}^{-3/2}$  whereas that for  $\text{SiBr}_4$  ( $347.9 \text{ g mol}^{-1}$ ) is  $18.0 \text{ J}^{1/2} \text{ cm}^{-3/2}$  in accord with the expected trend. These are nonpolar tetrahedral molecules that interact as a result of London dispersion forces. The much higher solubility parameters for  $\text{XeF}_2$  and  $\text{XeF}_4$  thus appear to be anomalous when compared to those for other nonpolar molecules.

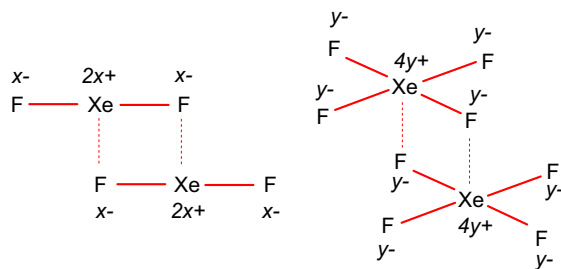
It must be remembered that although the *molecules* are nonpolar, the *bonds* are distinctly polar. Fluoride bridges are possible, but  $\text{XeF}_2$  and  $\text{XeF}_4$  are linear and square planar, respectively, in the solid and vapor states. Moreover, the solubility parameter for  $\text{XeF}_2$  is *higher* than that for  $\text{XeF}_4$ , which is contrary to the case when the molecular weights are considered. A possible explanation for the solubility parameters being higher than those of molecules such as  $\text{SnCl}_4$  and  $\text{SiBr}_4$  lies in the interaction of the molecules by means of *bond* dipoles that can be represented as shown in Figure 18.1.

However, the fact that the Xe atom in  $\text{XeF}_4$  has four fluorine atoms attached would lead to each Xe–F bond being less polar than would be the bonds in  $\text{XeF}_2$  where the polarization of Xe occurs in only two directions. Accordingly, the magnitude of the forces between molecules would be greater in  $\text{XeF}_2$  than it would be in  $\text{XeF}_4$ , in accord with the values of their solubility parameters and melting points.

### 18.3 REACTIONS OF XENON FLUORIDES

Much of the chemistry of the xenon fluorides centers on their reactions with aqueous solutions. Owing to its greater ionic character, the reaction of  $\text{XeF}_2$  with water is much slower than those of  $\text{XeF}_4$  and  $\text{XeF}_6$ . In fact,



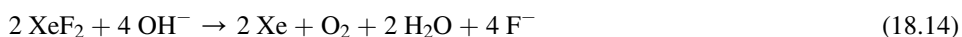


**FIGURE 18.1** Possible interactions between xenon difluoride and xenon tetrafluoride molecules as a result of bond dipoles. The magnitude of  $x$  would be larger than that of  $y$ , because the Xe atom is polarized in two directions (by two fluorine atoms) in the difluoride and in four directions (by four fluorine atoms) in the tetrafluoride.

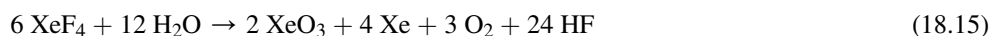
$\text{XeF}_2$  can be dissolved in water to produce relatively stable solutions. The slow hydrolysis reaction of  $\text{XeF}_2$  can be shown as follows:



In basic solution, the reaction is very rapid and is represented by the equation



The reaction of  $\text{XeF}_4$  with water is extremely vigorous and it occurs as a result of a disproportionation reaction that is similar to that of the halogens.



In this process, Xe(IV) reacts to produce Xe(VI) and Xe(0). The oxide,  $\text{XeO}_3$ , is a very dangerously explosive compound. Hydrolysis of  $\text{XeF}_6$  also produces this oxide as a product of the following reaction:



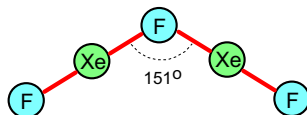
However, this reaction appears to take place in two steps as shown by the following equations:



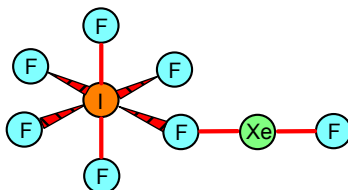
In Chapter 17, several reactions of interhalogens were shown in which cations were produced by a reaction with a strong fluoride ion acceptor (Lewis acid). One reaction of this type is the following:



The xenon fluorides undergo similar reactions with fluoride acceptors such as  $\text{SbF}_5$ ,  $\text{AsF}_5$ ,  $\text{TaF}_5$ ,  $\text{PtF}_5$ , etc. The reaction of  $\text{XeF}_2$  with  $\text{SbF}_5$  can produce  $\text{XeF}^+ \text{Sb}_2\text{F}_{11}^-$ ,  $\text{XeF}^+ \text{SbF}_6^-$ , or  $\text{Xe}_2\text{F}_3^+ \text{SbF}_6^-$  with  $\text{F}^-$  bridges between Xe centers in the cation. Thus,  $\text{Xe}_2\text{F}_3^+$  represents a fluoride ion bridging between two  $\text{XeF}^+$  ions as shown here.



In species where  $\text{XeF}^+$  and  $\text{MF}_6^-$  ions are present, the ions are associated by the formation of a structure in which there is a bridging  $\text{F}^-$  ion as shown in the structure that follows.



The  $\text{XeF}_3^+$  cation is generated when  $\text{XeF}_4$  reacts with  $\text{BiF}_5$  as shown by the equation



Solid  $\text{XeF}_6$  contains  $\text{XeF}_5^+$  ions that have fluoride ions bridging between them to give a cation that has the composition,  $\text{Xe}_2\text{F}_{11}^+$ . That cation is also generated when  $\text{XeF}_6$  reacts with pentafluorides such as  $\text{RuF}_5$ ,



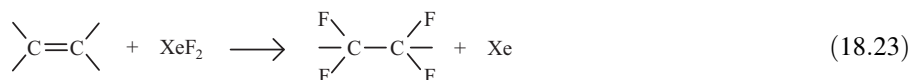
and the cation has the arrangement  $\text{F}_5\text{Xe}^+\cdots\text{F}^-\cdots\text{XeF}_5^+$ .

In some cases,  $\text{XeF}_6$  itself forms anions such as  $\text{XeF}_7^-$  and  $\text{XeF}_8^{2-}$  by virtue of its Lewis acidity. A general reaction can be written as

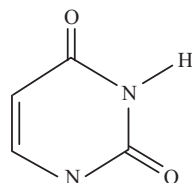


When heated, some compounds of this type undergo a reaction to produce  $\text{M}_2\text{XeF}_8$  and  $\text{XeF}_6$ .

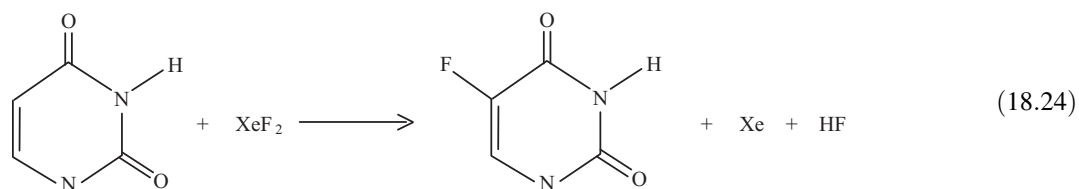
In chemical behavior similar to that of interhalogens, the xenon fluorides act as fluorinating agents in a wide variety of reactions. Xenon difluoride is a milder reagent than is the tetrafluoride or the hexafluoride. It readily reacts with olefins to add fluorine.



The structure of the uracil molecule can be shown as



A useful derivative of this compound is 5-fluorouracil that is used as a topical ointment in the treatment of some types of skin cancer and other skin diseases. One preparation of the compound is by the fluorination of uracil with  $\text{XeF}_2$ .



## 18.4 OXYFLUORIDES AND OXIDES

Xenon oxides are prepared from the fluorides. As has already been mentioned, hydrolysis of  $\text{XeF}_4$  and  $\text{XeF}_6$  leads to the formation of  $\text{XeO}_3$  by reactions that can be shown as

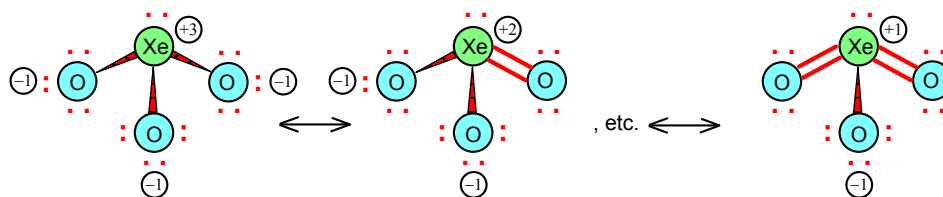


The latter reaction appears to involve two steps that can be described by the equations



These reactions are similar to the hydrolysis of  $\text{PCl}_5$  that produces  $\text{OPCl}_3$  when a limited amount of water is present except in that case the final product is an oxyhalide.

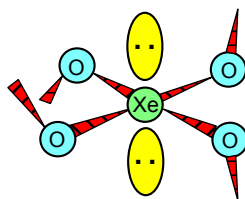
Solid  $\text{XeO}_3$  is a violently explosive white solid that has a very high positive heat of formation of approximately  $400 \text{ kJ mol}^{-1}$ . Because the four atoms in  $\text{XeO}_3$  have a total of 26 valence shell electrons, the predicted structure contains three bonds and an unshared pair of electrons surrounding the xenon atom. However, the resulting +3 formal charge on the xenon atom indicates that there should be contributions to the actual structure from resonance structures having multiple bonds between Xe and O. Therefore, resonance can be illustrated by the structures



More recently, Brock and Schrobilgen (2011) reported the synthesis of  $\text{XeO}_2$  by reaction of  $\text{XeF}_4$  with water at  $0^\circ\text{C}$ .



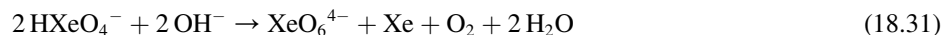
The compound degrades slowly at  $-78^\circ\text{C}$  and has an extended structure in which each xenon atom is surrounded by four oxygen atoms in a square planar arrangement as shown in the structure (adapted from Brock and Schrobilgen (2011)).



In the presence of  $\text{OH}^-$ , the reaction of  $\text{XeO}_3$  leads to the formation of a hydrogen xenate ( $\text{HXeO}_4^-$ ) ion.



Disproportionation of this unstable species in basic solutions occurs as represented by the equation

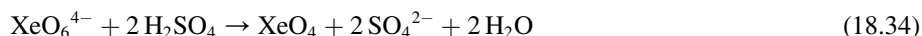


Solid perxenate ( $\text{XeO}_6^{4-}$ ) salts can be obtained that contain cations of Group IA and IIA metals. The  $\text{XeO}_6^{4-}$  ion has a very weak conjugate acid ( $\text{HXeO}_6^{3-}$ ) so the hydrolysis reactions



are extensive and the solutions are basic as well as very strong oxidizing agents.

In 1964, Selig and coworkers reported the synthesis of  $\text{XeO}_4$  by reaction of perxenate compounds with concentrated sulfuric acid according to the reaction



$\text{XeO}_4$  is a yellow solid with the oxygen atoms arranged in a tetrahedron around the xenon atom. It is relatively stable below  $-36^\circ\text{C}$  but is explosive above that temperature.

In reactions that are very similar to that between  $\text{PCl}_5$  and  $\text{P}_4\text{O}_{10}$ ,



xenon fluorides react with xenon oxides to produce oxyfluorides.



The introduction to the chemistry of the noble gases presented in this chapter is not exhaustive, but rather it is intended to give a survey of the important types of compounds and reactions that have been investigated. It should be apparent that the chemistry of these elements has come a long way since 1962.

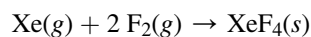
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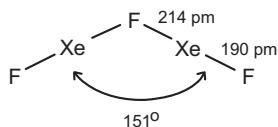
## PROBLEMS

- What attributes make  $\text{XeF}_2$  a desirable fluorinating agent for organic compounds?
- Which would you expect to be more stable,  $\text{NaXeF}_7$  or  $\text{CsXeF}_7$ ? Explain your answer.
- One of the uses of helium is as an inert atmosphere during welding, especially of aluminum. From a chemical point of view, why is helium a more appropriate choice for this use than would be nitrogen?
- Write complete balanced equations for the following.
  - $\text{XeF}_4 + \text{SF}_4 \rightarrow$
  - $\text{XeO}_3 + \text{OH}^- \rightarrow$
  - $\text{XeF}_2 + \text{OH}^- \rightarrow$
  - $\text{XeF}_2 + \text{P}_4 \rightarrow$
  - $\text{XeF}_4 + \text{PtF}_5 \rightarrow$
- If the  $\text{Kr}-\text{F}$  bond enthalpy is  $50 \text{ kJ mol}^{-1}$  and the  $\text{F}-\text{F}$  bond enthalpy is  $159 \text{ kJ mol}^{-1}$ , what would be the heat of formation of gaseous  $\text{KrF}_2$  from the gaseous elements?
- Draw structures for the following species.
  - $\text{SbF}_6^-$
  - $\text{XeO}_4$
  - $\text{XeF}_3^+$
  - $\text{XeF}_4$
  - $\text{XeF}_5^+$
- Discuss the role of multiple bonding in the  $\text{XeO}_4$  molecule.

8. Why is it possible for a molecule such as  $\text{XeF}_6$  or  $\text{N}_2\text{O}_5$  that exists as molecules in the gaseous state to take on a different nature in the solid state?
9. The  $\text{Xe}-\text{F}$  bond enthalpy is  $133 \text{ kJ mol}^{-1}$  and the  $\text{F}-\text{F}$  bond enthalpy is  $159 \text{ kJ mol}^{-1}$ . Calculate the heat of formation of gaseous  $\text{XeF}_2$  and  $\text{XeF}_4$ .
10. If the lattice energy of  $\text{XeF}_4(\text{s})$  is  $62 \text{ kJ mol}^{-1}$ , what would be the enthalpy change for the following reaction?



11. Suppose it would be possible to synthesize  $\text{ArF}_2(\text{g})$  even if it had a heat of formation of  $+100 \text{ kJ mol}^{-1}$ . What would the  $\text{Ar}-\text{F}$  bond enthalpy have to be to make the synthesis energetically feasible if the  $\text{F}-\text{F}$  bond enthalpy is  $159 \text{ kJ mol}^{-1}$ ?
12. Given the fact that the  $\text{Xe}-\text{F}$  bond enthalpy is  $133 \text{ kJ mol}^{-1}$  and the  $\text{Kr}-\text{F}$  bond enthalpy is only  $50 \text{ kJ mol}^{-1}$ , comment on the possibility of preparing  $\text{ArF}_2$  and  $\text{NeF}_2$ .
13. Explain why the bond lengths in the  $\text{Xe}_2\text{F}_3^+$  ion are unequal as shown in the structure



# The Transition Metals

The elements that are positioned in the periodic table between those in the first two groups and those in the last six groups (the so-called main groups) are known as the transition metals. These metals, except the last two in each series, are characterized by having  $d$  orbitals that are only partially filled. Because there are three such series for which the partially filled  $d$  orbitals are the  $3d$ ,  $4d$ , and  $5d$  orbitals, the series are referred to as the first-, second-, and third-row transition series, respectively. The elements in these groups are usually designated in different ways as either Groups 3–12 (when the groups are simply numbered as 1 through 18 when going left to right in the periodic table) or as the “B” group elements.

Although some of the metals in the second and third transition series are certainly important, the first-row metals constitute the extremely significant and useful structural metals. In addition to the metals themselves, an enormous number of alloys containing these metals are widely used as are many compounds of the metals. Of all the compounds of the transition metals, titanium dioxide,  $\text{TiO}_2$ , is produced in the largest quantity (approximately 2 billion lb/yr) because of its extensive use in paints.

For several reasons, the majority of the discussion in this chapter will deal mainly with the first-row metals, but many properties of the second- and third-row metals are similar to those of the first row that have the same  $d^n$  configuration. The lanthanides constitute a series within a transition series so they will also be discussed briefly. Several additional aspects of the chemistry of the transition metals in all three series will be discussed in Chapters 19–22 in dealing with coordination compounds and organometallic chemistry.

## 19.1 THE METALS

A summary of general information regarding the occurrence, properties, and uses of the first-row transition metals is shown in [Table 19.1](#). Most of the second- and third-row transition metals are found as minor constituents in ores containing other metals. Consequently, we will not enumerate the sources, minerals, or the processes by which these metals are obtained. However, some of their most important properties are presented in [Table 19.2](#).

### 19.1.1 Structures of Metals

As was described in Chapter 4, the structures of metals are determined by the various ways to efficiently pack spheres. With regard to structures, efficiency refers to minimizing the amount of free (empty) space in the structure and maximizing the number of atoms that are in simultaneous contact (the extent of metallic bonding). Three of the most common ways to arrange spherical atoms are shown in [Figure 19.1](#).

The simplest but least common structure known for metals is the simple cubic structure shown in [Figure 19.1\(a\)](#). It consists of atoms located on the corners of a cube. Because eight cubes come together at each corner when an extended array is constructed, only one-eighth of each atom on the corners belongs to any one cube. Thus, there is  $8(1/8) = 1$  atom per unit cell. In this structure, each atom is surrounded by six other metal atoms (nearest neighbors), so the coordination number is 6. A calculation of the percent of empty space in the unit cell is easily performed by assuming that the spheres on the corners of the cube touch at the midpoint of an edge. The volume of the cube is  $V_c = e^3$ , where  $e$  is the length of an edge of the cube. The radius of an atom is  $e/2$  and its volume is  $V_a = (4/3)\pi(e/2)^3 = 0.524 e^3$ . Thus, only 52.4% of the volume of the unit cell is occupied by the single atom constituting the unit cell. Consequently, there is 47.6% free space in the simple cubic structure, which means that it is not a very efficient packing model. Moreover, each atom has only six nearest neighbors so this packing model does not maximize the number of bonds between atoms.

A somewhat more complex structure is the body-centered cubic (*bcc*) structure that has eight atoms on the corners of a cube and one atom located in the center of the cube (see [Figure 19.1\(b\)](#)) for a total of two atoms per unit cell. In this structure, each atom on a corner is in touch with the one in the center of the cube. As a result, a diagonal through the cube is equal to four times the radius (one diameter plus two times the radius) of an atom. The fraction of free space can be

**TABLE 19.1** The First-Row Transition Metals

	Minerals	Composition	Mineral Sources <sup>a</sup>	Process	Element Properties <sup>b</sup>	Uses
Sc	<i>Thortvetite</i>	37–42% Sc <sub>2</sub> O <sub>3</sub>	Norway, Madagascar	Reduction of ScF <sub>3</sub> with carbon or Zn/Mg alloy	<i>hcp</i> , 2.99, m.p. 1541 °C, b.p. 2836 °C, 160, 1.3	Semiconductor, isotope studies
	<i>Wolframite</i>	(Fe,Mn)WO <sub>4</sub>				
Ti	<i>Ilmenite</i>	FeTiO <sub>3</sub>	Russia, India, US, Canada, Sweden, West Africa, Australia	Reduction of TiCl <sub>4</sub> with Mg or Na	<i>hcp</i> , 4.5, m.p. 1660 °C, b.p. 3287 °C, 148, 1.5	Alloys, machinery aircraft, missiles
	<i>Rutile</i>	TiO <sub>2</sub>				
	<i>Titanite</i>	CaTiSiO <sub>5</sub>				
	<i>Perovskite</i>	CaTiO <sub>3</sub>				
V	<i>Patronite</i>	Approx. VS <sub>4</sub>	CO, UT, NM, AZ, SD, Mexico, Peru, Zaire, S. Africa, Russia	Reduction of V <sub>2</sub> O <sub>5</sub> with Ca	<i>bcc</i> , 6.11, m.p. 1890 °C, b.p. 3380 °C, 131, 1.6	Alloy steels, catalysts, X-ray target
	<i>Roscoelite</i>	K <sub>2</sub> V <sub>4</sub> Al <sub>2</sub> Si <sub>6</sub> O <sub>20</sub> (OH) <sub>4</sub>				
	<i>Carnotite</i>	K <sub>2</sub> (UO <sub>2</sub> ) <sub>2</sub> (VO <sub>4</sub> ) <sub>2</sub> ·3 H <sub>2</sub> O				
	<i>Vanadinite</i>	Pb <sub>3</sub> Cl(VO <sub>4</sub> ) <sub>3</sub>				
Cr	<i>Chromite</i>	FeCr <sub>2</sub> O <sub>4</sub>	Russia, Philippines, Cuba, Turkey, S. Africa	Reduction of FeCr <sub>2</sub> O <sub>4</sub> with Al or C	<i>bcc</i> , 7.19, m.p. 1900 °C, b.p. 2672 °C	Stainless steels, plating, pigments
Mn	<i>Pyrolucite</i>	MnO <sub>2</sub>	Germany, Australia, Canada, India, Europe, US	Reduction of oxide with Al or C	<i>fcc</i> , 7.44, m.p. 1244 °C, b.p. 1962 °C, 112, 1.5	Steel alloys, chemicals
	<i>Manganite</i>	MnO(OH)				
	<i>Psilomelane</i>	BaMn <sup>2+</sup> Mn <sup>4+</sup> <sub>8</sub> O <sub>16</sub> (OH) <sub>4</sub>				
	<i>Rhodochrosite</i>	MnCO <sub>3</sub> (some Fe, Ca)				
Fe	<i>Hematite</i>	Fe <sub>2</sub> O <sub>3</sub> (impure)	Labrador, Yukon, S. America, Europe, US	Reduction of oxide with C	<i>bcc</i> , 7.87, m.p. 1535 °C b.p. 2750 °C, 124, 1.8	Steel making, powder metallurgy, magnets, rails, structures
	<i>Limonite</i>	FeO(OH)·nH <sub>2</sub> O				
	<i>Magnetite</i>	Fe <sub>3</sub> O <sub>4</sub> (some Ti, Mn)				
	<i>Siderite</i>	FeCO <sub>3</sub>				

Co	<i>Skutterudite</i>	$(\text{Co,Ni,Fe})\text{As}_3$	Canada, Zaire, Sweden, N. Africa	Roast ore/reduce oxide with Al or electrolysis of solutions	<i>hcp</i> , 8.90, m.p. 1943 °C, b.p. 2672 °C, 125, 1.9	Alloys, magnets, chemicals, jet engine parts, catalysts
	<i>Cobaltite</i>	$\text{CoAsS}$				
	<i>Erythrite</i>	$(\text{Co})_3(\text{AsO}_4)_2 \cdot 8 \text{H}_2\text{O}$				
Ni	<i>Pentlandite</i>	$(\text{Fe,Ni})_x\text{S}_y$	Canada, S. Africa, Russia, New Caledonia	Leaching with $\text{NH}_3$	<i>fcc</i> , 8.91, m.p. 1453 °C, b.p. 2732 °C	Alloys (steel, brass), batteries, catalysts
	<i>Garnierite</i>	$(\text{Ni,Mg})_3\text{Si}_2\text{O}_5(\text{OH})_4$				
Cu	<i>Cuprite</i>	$\text{Cu}_2\text{O}$	MI, AZ, UT, MT, MN, NV, Mexico, Chile, Peru, Canada, Africa, Russia	Roast, concentrate, flotation, leaching, reduce, electrolytic refining	<i>fcc</i> , 8.94, m.p. 1083 °C, b.p. 2567 °C, 128, 1.9	Wiring, switches, alloys (brass, bronze, Monel), plumbing
	<i>Bornite</i>	$\text{Cu}_6\text{FeS}_4$				
	<i>Chalcocite</i>	$\text{Cu}_2\text{S}$				
	<i>Chalcopyrite</i>	$\text{CuFeS}_2$				
	<i>Covellite</i>	$\text{CuS}$				
	<i>Malachite</i>	$\text{CuCO}_3 \cdot \text{Cu}(\text{OH})_2$				
	<i>Azurite</i>	$\text{Cu}_3(\text{CO}_3)_2(\text{OH})_2$				
Zn	<i>Zinc blende</i>	$\text{ZnS}$	MO, KS, CO, OK, ID, MI, UT, Australia, Canada, Mexico, Europe	Roast to oxide, reduce with C and distill, also hydrometal	<i>hcp</i> , 7.14, m.p. 420 °C, b.p. 907 °C, 133, 1.6	Alloys (brass, bronze), coatings, batteries, cable shielding
	<i>Hemimorphite</i>	$\text{Zn}_4\text{Si}_2\text{O}_7(\text{OH})_2 \cdot 2 \text{H}_2\text{O}$				
	<i>Smithsonite</i>	$\text{ZnCO}_3$				
	<i>Hydrozincite</i>	$\text{Zn}_5(\text{OH})_6(\text{CO}_3)_2$				
	<i>Sphalerite</i>	$\text{ZnS}$				
	<i>Willemite</i>	$\text{Zn}_2\text{SiO}_4$				
	<i>Wurtzite</i>	$\text{ZnS}$				

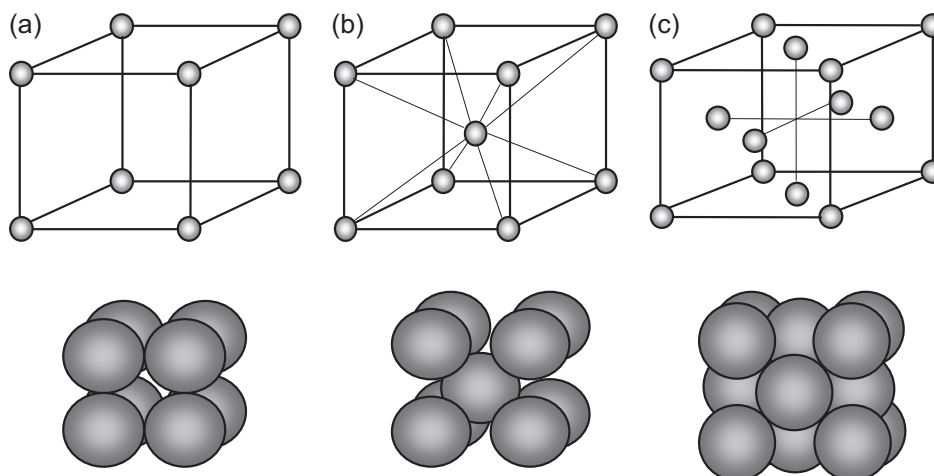
<sup>a</sup>Not all of the minerals occur in each country.

<sup>b</sup>Information included is crystal structure, density (in  $\text{g cm}^{-3}$ ), melting point, boiling point, atomic radius (in pm), and electronegativity, in that order.



**TABLE 19.2** Properties of Second- and Third-Row Transition Metals

Transition Metals of the Second Row										
	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd
m.p., °C	1522	1852	2468	2617	2172	2310	1966	1552	962	321
b.p., °C	3338	4377	4742	4612	4877	3900	3727	3140	2212	765
Crystal structure	<i>hcp</i>	<i>hcp</i>	<i>bcc</i>	<i>bcc</i>	<i>fcc</i>	<i>hcp</i>	<i>fcc</i>	<i>fcc</i>	<i>fcc</i>	<i>hcp</i>
Density, g cm <sup>-3</sup>	4.47	6.51	8.57	10.2	11.5	12.4	12.4	12.0	10.5	8.69
Atomic radius, pm	182	162	143	136	136	134	134	138	144	149
Electronegativity	1.2	1.4	1.6	1.8	1.9	2.2	2.2	2.2	1.9	1.7
Transition Metals of the Third Row										
	La	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg
m.p., °C	921	2230	2996	3407	3180	3054	2410	1772	1064	-39
b.p., °C	3430	5197	5425	5657	5627	5027	4130	3827	2807	357
Crystal structure	<i>hcp</i>	<i>hcp</i>	<i>bcc</i>	<i>bcc</i>	<i>hcp</i>	<i>hcp</i>	<i>fcc</i>	<i>fcc</i>	<i>fcc</i>	-
Density, g cm <sup>-3</sup>	6.14	13.3	16.7	19.3	21.0	22.6	22.6	21.4	19.3	13.6
Atomic radius, pm	189	156	143	137	137	135	136	139	144	155
Electronegativity	1.0	1.3	1.5	1.7	1.9	2.2	2.2	2.2	2.4	1.9

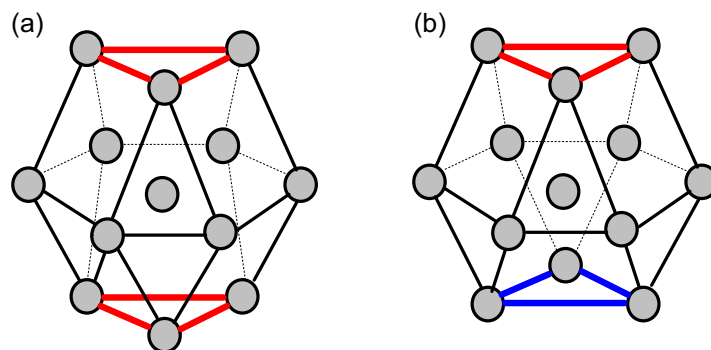
**FIGURE 19.1** The (a) simple cubic (*sc*), (b) body-centered cubic (*bcc*), and (c) face-centered cubic (*fcc*) closest packing models. The spheres are considered to be touching but in the upper structures they are shown smaller with spaces between them for clarity.

calculated as follows. The length of the diagonal through the cube is  $3^{1/2}e$  or  $4r$  where  $r$  is the radius of an atom. Accordingly, the atomic radius can be expressed as

$$\text{radius} = r = \left( \frac{e(3)^{1/2}}{4} \right)^3 \quad (19.1)$$

where  $e$  is the length of the edge of the cube. Thus, the total volume of the two atoms in the *bcc* unit cell can be represented as

$$\text{Volume} = V = 2 \left[ \frac{4\pi}{3} \left( \frac{e(3)^{1/2}}{4} \right)^3 \right] = 0.680e^3$$



**FIGURE 19.2** The (a) hexagonal and (b) face-centered cubic closest packing models. The spheres are considered to be touching but are shown smaller for clarity.

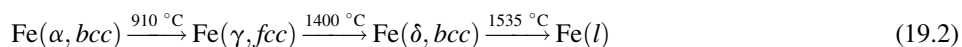
With 68.0% of the volume of the cell being occupied by the two atoms, the fraction of free space in this structure is 32.0%. Therefore, it represents a more efficient packing model than does the simple cubic model. Also, each atom is surrounded by eight nearest neighbors in the *bcc* structure, and there are two atoms per unit cell, which means it is preferable to the simple cubic structure on this basis also.

A structure that is exhibited by many metals is the face-centered cubic (*fcc*) structure which is also referred to as *cubic closest packing*. This structure has a unit cell consisting of one atom on each corner of a cube as well as one atom on each face that is shared between the two cubes joined at that face (see Figure 19.1(c)). Because a cube has six faces and eight corners, the number of atoms per unit cell is  $6(1/2) + 8(1/8) = 4$ . Each atom is surrounded by 12 nearest neighbors, and although the calculation will not be shown, the free space is only 26% of the volume of the cubic cell. Thus, the *fcc* (cubic closest packing) structure is the most efficient of the three cubic structures described.

A fourth structure that is also frequently exhibited by metals is the hexagonal closest packing (*hcp*) model shown in Figure 19.2(a). In the *hcp* structure, each atom is surrounded by six others in a hexagonal arrangement. The layers lying above and below this layer each allow three additional atoms to be brought into contact with each atom in the layer shown as the “hexagonal” layer. This structure results in a coordination number of 12 for each atom.

As in the case of the *fcc* structure, the *hcp* structure also results in 26% free space. The fraction of *occupied* space in the unit cell is sometimes called the *packing factor*. Thus, both the *fcc* and *hcp* models have packing factors equal to 0.74. The difference between the *hcp* and *fcc* models is that in *hcp* the layers repeat after two layers of atoms whereas in *fcc* they repeat after three layers. When the arrangements of atoms in the layers are designated as A, B, and C, these arrangements can be described as ...ABABAB... and ...ABCABCABC..., respectively. In each model, one atom is surrounded by six others in a hexagonal plane and three others from each of the layers lying above and below. Therefore, a total of 12 atoms touch the one shown as the central atom in the hexagonal plane with three coming from above and three from below.

Each of the structures shown in Figure 19.2 has a coordination number of 12. If the layers above and below the hexagonal plane are identical (the atoms are aligned), the packing model is hexagonal closest packing. If the layers above and below are not identical (the atoms are staggered), the packing model is cubic closest packing. Both models involve the same coordination number and the same percent of free space, so these arrangements have approximately equal energy. It should not be surprising that a sizable number of cases are known in which a metal can be changed from one of these structural types to the other. Generally, these phase transitions are brought about by heat or a change in pressure. Consequently, iron exists in at least four forms or phases, and cobalt and nickel exist in at least two forms each. For iron, some of the transformations that occur at 910 °C and 1400 °C can be shown as follows (iron melts at 1535 °C).



A change in structure occurs in solid cobalt at 417 °C and in titanium at 883 °C.



These are only a few of the phase changes that are known to occur in transition metals.

Table 19.1 lists several uses of the first-row transition metals. Because they are, in general, common uses, only a very brief survey will be presented here. As a result of its resistance to corrosion and the fact that it has high strength combined with light weight, titanium and its alloys are of great utility in aircraft construction where strong but lightweight objects are essential. Vanadium and aluminum are alloyed with titanium to produce alloys that are stronger than titanium itself. One of the common alloys contains 6% aluminum and 4% vanadium.

Although the major use of vanadium, manganese, and cobalt is as alloying metals with iron in making specialty steels, chromium is widely used as a plating metal because it is highly resistant to corrosion. Of even greater importance is its use in stainless steels that will be addressed in the next section.

The metallurgy of iron-based alloys is a vast and complex field, and more iron is used than all other metals combined. Among the many forms of iron are cast iron, wrought iron, and the myriad special steels that contain other metals as well as carbon and other main group elements.

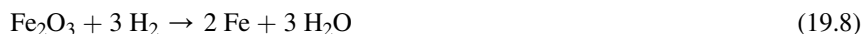
The importance of iron and its alloys is impossible to overemphasize. They constitute over 90% of all the metals utilized, and the number of alloy compositions containing iron is enormous. Even “iron” is not the pure metal but rather it contains some carbon because of its use in the reduction process that can be shown in simplified form as



In some cases, methane is injected into the blast furnace with preheated air, and hydrogen is produced.



The hydrogen can then act as a reducing agent in addition to the carbon monoxide.



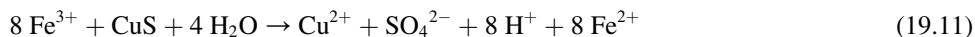
One form of cast iron, sometimes known as white iron, is a material that contains iron carbide,  $\text{Fe}_3\text{C}$ , which is also known as *cementite*. Gray iron has a lower content of carbon but a higher silicon content that results in carbon being present in a solid graphite phase instead of  $\text{Fe}_3\text{C}$ . If a small amount (0.05%) of magnesium is added, the result is a form of iron known as ductile iron that is much more malleable and strong. Heat treating white iron increases its malleability by changing its structure. This form is closely related to wrought iron. Although referred to as “iron,” all of these forms are actually alloys containing iron and carbon (approximately 2–5%). A very brief introduction to steels will be given in the next section. As can be seen from this brief introduction, the characteristics of “iron” are greatly dependent on the nature and amount of other substances.

Copper forms alloys with many other metals including tin, zinc, nickel, and aluminum, and because of its conductivity is widely used in electrical applications. It is found primarily as the ore *chalcopyrite*,  $\text{CuFeS}_2$  and is reduced by roasting in air.



The  $\text{SO}_2$  produced in this manner is recovered and further reacted to form sulfuric acid.

Copper can also be obtained from mining tailings and low-grade ores by the use of the bacterium *thiobacillus ferrooxidans*. This bacterium thrives in an acidic environment and decomposes iron sulfides and oxidizes  $\text{Fe}^{2+}$  to  $\text{Fe}^{3+}$ . The  $\text{Fe}^{3+}$  then oxidizes the sulfide ion in copper sulfides to form copper sulfate, which is soluble.



The  $\text{Cu}^{2+}$  ion is reduced by reaction with scrap iron through the reaction



to form solid copper.

One form of nickel is an important catalyst known as *Raney nickel* that is prepared by the reduction of  $\text{NiO}$  with hydrogen. Nickel also is used in several alloys that have wide application. For example, *Monel* is a type of alloy that contains nickel and copper in a ratio of about 2:1. It is frequently used in making bathroom fixtures.

Zinc and its alloys are widely used because they have low melting points and can be easily cast. Therefore, a wide range of objects are made from zinc and its alloys. Zinc is widely applied to iron and steel as a protective coating by the process known as *galvanization* that consists of coating an iron object with a thin layer of zinc. Relative to iron, zinc is an anode so it is preferentially oxidized. If the coating is broken, the zinc continues to corrode rather than the iron object. When iron is coated with a less reactive metal such as tin, a break in the coating causes the more easily oxidized iron to be corroded at an accelerated rate.

Some second- and third-row transition metals are, for good reason, known as precious metals. As this is written, these include silver (\$17/oz), palladium, rhodium, iridium, osmium, gold (\$1200/oz), and platinum (\$1200/oz). The other metals in this category range from a few hundred dollars per ounce to well over \$2000 per ounce.

### 19.1.2 Alloys

As we have seen, the characteristics of metals that permit them to function as enormously versatile construction materials are their ductility, malleability, and strength. The last of these characteristics needs no explanation at this time, but the first two are related to the ability of the metal to be fabricated into a desired shape. Metals vary widely in these characteristics, and a metal or alloy that is well suited to one use may be entirely unsatisfactory for another. Addressing this branch of applied science is beyond the scope of this book, but a book on materials science provides a great deal of information that is relevant for students in inorganic chemistry.

As important as the transition metals are, alloys greatly extend their applicability. In this section, we will describe some of the major factors that are important in the behavior of alloys. The study of alloys is a vast area of applied science so in the interest of brevity the emphasis will be primarily on the behavior of alloys of copper and iron. However, many of the principles that determine the behavior of one metal are applicable to the behavior of others. Alloys of some of the non-transition metals (lead, antimony, tin, etc.) have been described in previous chapters.

Alloys are classified broadly in two categories, *single-phase* alloys and *multiple-phase* alloys. A *phase* is characterized by having a homogeneous composition on a macroscopic scale, a uniform structure, and a distinct interface with any other phase present. The coexistence of ice, liquid water, and water vapor meets the criteria of composition and structure, and distinct boundaries exist between the states so there are three phases present.

When liquid metals are combined, there is usually some limit to the solubility of one metal in another. An exception to this is the liquid mixture of copper and nickel that forms a solution of any composition ranging from pure copper to pure nickel. When the mixture is cooled, a solid results that has a random distribution of both types of atoms in an *fcc* structure. This single solid phase thus constitutes a solid solution of the two metals so it meets the criteria for a single-phase alloy.

Alloys of copper and zinc can be obtained by combining the molten metals. However, zinc is soluble in copper up to only about 40% (of the total mass). When the content of a copper/zinc alloy contains less than 40% zinc, cooling the liquid mixture results in the formation of a solid solution in which Zn and Cu atoms are uniformly distributed in an *fcc* lattice. When the mixture contains more than 40% zinc, cooling the liquid mixture results in the formation of a compound having the composition CuZn. The solid alloy consists of two phases, one of which is the compound CuZn and the other is a solid solution that contains Cu with approximately 40% Zn dissolved in it. This type of alloy is known as a two-phase alloy, but many alloys contain more than three phases (multiple phase alloys).

The formation of solid solutions of metals is one way to change the properties (generally to increase strength) of the metals. Strengthening metals in this way is known as solid solution strengthening. The ability of two metals to form a solid solution can be predicted by a set of rules known as the *Hume-Rothery rules* that can be stated as follows.

1. The atomic radii of the two kinds of atoms must be similar (within about 15%) so that lattice strain will not be excessive.
2. The crystal structure of the two metals must be identical.
3. In order to minimize the tendency for the metals to form compounds, identical valences and approximately equal electronegativities are required.

However, it must be mentioned that these guidelines are not always successful for predicting solubility of metals.

When a second metal is dissolved in a host metal and the mass is cooled to produce a solid solution, the solution has increased strength compared to the host metal. This occurs because in a regular lattice consisting of identical atoms it is relatively easy to move the atoms away from each other. There is less restriction on mobility because the electron sharing between atoms is equal. However when zinc or nickel is dissolved in copper, the alloy is strengthened, and the degree of strengthening is approximately a linear function of the fraction of the added metal. Adding a second metal distorts the lattice at the sites where the atoms of that substituted metal reside, and this restricts atomic motions and

strengthens the metal. However, the atomic radius of copper is 128 pm and those of zinc and nickel are 133 and 124 pm, respectively. The atomic radius of tin is 151 pm so if tin were dissolved in copper (assuming that its limited solubility is not exceeded), there would be a greater strengthening effect because there is a greater disparity between the sizes of the host and guest metal atoms. This is exactly the observed effect, and bronze (Cu/Sn) is stronger than brass (Cu/Zn). The atomic radius of beryllium is 114 pm so adding beryllium (within the solubility limit) also causes a great strengthening of copper. In fact, the addition of atoms having a smaller atomic radius than that of copper causes a greater effect than adding metals whose atoms are larger than copper even when the same absolute difference in size exists. In both cases, the degree of strengthening is approximately a linear function of the weight percent of the metal added to copper. An interesting observation regarding the effect of having some different atoms in a lattice is illustrated by the alloy Monel. Nickel is harder than copper, but when the alloy containing the two metals is made, it is harder than nickel.

No single volume could contain the complete description of the composition, properties, and structures of ferrous alloys. Further, the effect of heat treatment and other methods of changing the properties of alloys constitutes an entire science unto itself. Accordingly, the description of ferrous metallurgy given here will be only a brief overview of this important area.

Steels constitute a wide range of iron-based alloys. General types include carbon steels (containing from 0.5% to 2.0% carbon) and only small amounts of other metals (generally less than 3–4%). Other metals in the alloys may include nickel, manganese, molybdenum, chromium, or vanadium in varying amounts. These ingredients are included to produce a steel having the desired characteristics. The properties of the steel are determined by the composition as well as the heat treatment methods employed.

If the total amount of other metals present in iron exceeds about 5%, the alloy is sometimes called a *high-alloy* steel. Most stainless steels are in this category because the chromium content is between 10% and 25%, and some types also contain 4–20% nickel. Stainless steels, so-called because of their resistance to corrosion, are of several types. The form of iron having the *fcc* structure is known as  $\gamma$ -Fe or *austenite*. Therefore, one type of stainless steel (that also contains nickel) is known as *austenitic* stainless steel because it has the austenite (*fcc*) structure. *Martensitic* stainless steels have a structure that contains a body-centered tetragonal arrangement that results from rapidly quenching the austenite structure. In addition to these types, *ferritic* stainless steel has a *bcc* structure, but it does not contain nickel. In addition to the stainless steels, a large number of alloys known as tool steels are important. As the name implies, these are special alloys that are used to make tools for cutting, drilling, and fabricating operations. These alloys commonly include some or all of the following elements in varying amounts: Cr, Mn, Mo, Ni, W, V, Co, C, and Si. In many cases, the alloys are engineered to have the desired properties of resistance to impact, heat, abrasion, corrosion, or thermal stress. Treatment of the steel having the desired composition can alter the structure of the metal so that the desired properties are optimized. As a result, the manufacture of steel is a complex process that involves many variables. The manufacture of special steels is an important area of metallurgy that may not be fully appreciated when driving an automobile that has dozens of the alloys used in its construction.

Alloys that retain high strength at high temperatures ( $>1000\text{ }^{\circ}\text{C}$  in some cases) are known as *superalloys*. Some of these materials are also highly resistant to corrosion (oxidation). These alloys are difficult to make and contain metals that are not readily available, and as a result, are expensive. They are used in situations where the conditions of service make them essential, e.g., aircraft engines where certain designs require as much as 50% by weight of these special alloys.

The designation of an alloy as a superalloy is based on its strength at high temperature. In the application of these alloys in fabricating a gas turbine, this is important because the efficiency of a turbine is greater at high temperature. Frequently, it is possible to study the behavior of an object for an extended period of time as it is subjected to a stress that is not sufficient to cause failure of the object. Even though the object may not break, it may elongate due to stretching of the metal. Movement of the metal under stress is called *creep*. Not only do the superalloys have high strength at elevated temperatures, but also they are resistant to creep which makes them desirable for many uses.

Generally, superalloys are given special names, and a few of the more common alloys of this type and their compositions are described in [Table 19.3](#). Because several of the superalloys contain very little iron, they are closely related to the nonferrous alloys. A discussion of this important area is beyond the scope of this book but the interested reader should consult the references at the end of this chapter.

Some of the second- and third-row transition metals possess many of the desirable properties of superalloys. They maintain their strength at high temperatures, but they are somewhat reactive with oxygen under these conditions. These metals are known as *refractory* metals, and they include niobium, molybdenum, tantalum, tungsten, and rhenium.

**TABLE 19.3** Some of the Superalloys

Name	Composition, Percent by Weight
16-25-6	Fe 50.7; Ni 25; Cr 16; Mo 6; Mn 1.35; C 0.06
Haynes 25	Co 50; Cr 20; W 15; Ni 10; Fe 3; Mn 1.5; C 0.1
Hastelloy B	Ni 63; Mo 28; Fe 5; Co 2.5; Cr 1; C 0.05
Inconel 600	Ni 76; Cr 15.5; Fe 8.0; C 0.08
Astrolloy	Ni 56.5; Cr 15; Co 15; Mo 5.5; Al 4.4; Ti 3.5; C 0.6; Fe < 0.3
Udimet 500	Ni 48; Cr 19; Co 19; Mo 4; Fe 4; Ti 3; Al 3; C 0.08

## 19.2 OXIDES

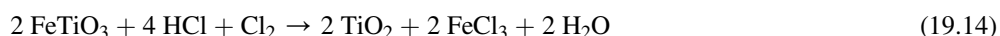
The oxides of the transition metals frequently show some degree of nonstoichiometric character. Most of the metals form more than one oxide so there is some tendency for oxidation to give a mixture of the oxides. When bulk metals react with oxygen, a layer of oxide forms through which additional oxygen or the metal must diffuse for the reaction to continue. Thus, there is a complex reaction between the metal, lower oxides, and higher oxides so that the reaction may not result in a single product of exact stoichiometry (see Chapter 5).

Because the chemistry of scandium is relatively straightforward, it will be discussed only briefly. Scandium has the  $4s^2 3d^1$  electron configuration so the usual compounds contain Sc(III). The only oxide that it forms is  $\text{Sc}_2\text{O}_3$ , and it is not a widely used compound. However, scandium is becoming increasingly important in fabricating metal objects that must have high strength and low weight. One such application is in the production of sports equipment and frames for lightweight firearms.

A very important oxide of titanium is  $\text{TiO}_2$  which is also the composition of one of its important ores, *rutile* (see Figure 4.6). As mentioned earlier,  $\text{TiO}_2$  is used in very large amounts as a paint pigment, partially due to its bright white color and its opacity. These characteristics give it good covering power, and it is a desirable pigment because it is essentially nontoxic. Reduction of  $\text{TiO}_2$  with aluminum gives the metal.



As a preparative technique,  $\text{TiO}_2$  can be obtained from  $\text{FeTiO}_3$ , *ilmenite*, by the reaction



It is an acidic oxide that reacts with basic oxides to form titanates ( $\text{TiO}_3^{2-}$ ) in a type of reaction that has been shown in other instances.

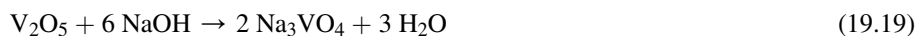
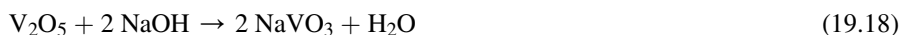


The mineral  $\text{CaTiO}_3$  is *perovskite*, and it is the compound that gives that crystal structure its name (see Figure 4.8).

Several oxides of vanadium are known with the simplest being VO although it is of little practical significance. The oxide  $\text{V}_2\text{O}_3$  can be obtained by the partial reduction of  $\text{V}_2\text{O}_5$  with hydrogen as is illustrated by the equation



The acidic character of  $\text{V}_2\text{O}_5$  is illustrated by its ability to form a series of oxyanions known as vanadates as shown in the following reactions:



The species  $\text{VO}_3^-$ ,  $\text{VO}_4^{3-}$ , and  $\text{V}_2\text{O}_7^{4-}$  are analogous to the phosphorus-containing meta-, ortho-, and pyrophosphate anions, respectively (see Chapter 14). In the last two of these, the structures contain the vanadium ion surrounded by

four oxide ions in a tetrahedral arrangement. This behavior emphasizes the nature of vanadium as a Group VB element. As in the case of the phosphate ion,  $\text{VO}_4^{3-}$  is a base that undergoes hydrolysis to give a basic solution.



In addition to the  $\text{V}_2\text{O}_7^{4-}$  species, other polyvanadates that contain large clusters can be obtained by adjusting the pH of the solutions. For example, the  $\text{V}_{10}\text{O}_{28}^{6-}$  ion contains  $\text{VO}_6$  octahedra in which an edge is shared. There are many other similarities to phosphorus chemistry including the formation of mixed oxyhalides of the type  $\text{OVX}_3$  (where  $\text{X} = \text{F}, \text{Cl},$  or  $\text{Br}$ ) and  $\text{VO}_2\text{X}$  (which is more correctly shown as  $\text{XVO}_2$ ).

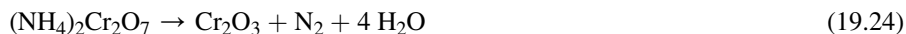
Chromium forms at least four oxides that represent the oxidation states of +2, +3, +4, and +6. The first of these is  $\text{CrO}$  that can be prepared by the reaction using an amalgam of  $\text{Cr}/\text{Hg}$ .



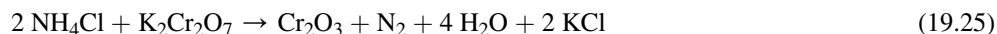
Although  $\text{CrO}$  is of little importance,  $\text{Cr}_2\text{O}_3$  is very important because of its ability to function as a catalyst for several reactions. This oxide can be obtained by several reactions that include oxidation of the metal,



and the decomposition of ammonium dichromate,



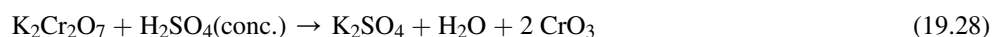
A mixture of  $\text{NH}_4\text{Cl}$  and  $\text{K}_2\text{Cr}_2\text{O}_7$  reacts in much the same way as does  $(\text{NH}_4)_2\text{Cr}_2\text{O}_7$  to produce  $\text{Cr}_2\text{O}_3$ .



This reaction emphasizes the acidic character of  $\text{NH}_4\text{Cl}$  as was also described in Section 6.1.3. The green color of  $\text{Cr}_2\text{O}_3$  has given rise to its being used as a pigment for many years. When the oxide is heated strongly, it is converted to a form that is unreactive toward acids, but it is amphoteric in its behavior as shown in the following equations.



Given the similar size of  $\text{Ti}^{4+}$  and  $\text{Cr}^{4+}$ , it should not be surprising that  $\text{CrO}_2$  has the rutile structure. However,  $\text{CrO}_2$  is a magnetic oxide that is used in magnetic devices and tapes. The oxide  $\text{CrO}_3$  is a strong oxidizing agent because it contains chromium in the +6 oxidation state. It results from the reaction of  $\text{K}_2\text{Cr}_2\text{O}_7$  with concentrated sulfuric acid.



Chromium(VI) oxide is a strongly acidic oxide that produces  $\text{H}_2\text{CrO}_4$  in aqueous solutions.



$\text{CrO}_3$  is such a strong oxidizing agent that it causes combustion of many organic materials. When heated to temperatures above  $250^\circ\text{C}$ , it decomposes as illustrated by the equation

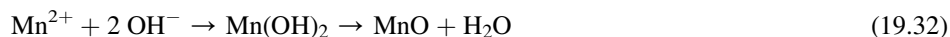


The principal ore of chromium is *chromite*,  $\text{Fe}(\text{CrO}_2)_2$ , which can also be written as  $\text{FeO} \cdot \text{Cr}_2\text{O}_3$ . Chromium(II) is a reducing agent that is easily oxidized to  $\text{Cr}^{3+}$ . In aqueous solutions,  $\text{Cr}^{2+}$  is present as the aqua complex  $[\text{Cr}(\text{H}_2\text{O})_6]^{2+}$  that is a beautiful blue color. A solution containing  $\text{Cr}^{2+}$  can be prepared by reducing  $\text{Cr}^{3+}$  or  $\text{Cr}_2\text{O}_7^{2-}$  with zinc in hydrochloric acid. Passing that solution into one containing acetate ions results in the formation of a brick red precipitate of  $\text{Cr}(\text{C}_2\text{H}_3\text{O}_2)_2$ . One unusual aspect of this compound is that it is one of a very small number of acetates that is insoluble. The chemistry of  $\text{Cr}^{3+}$  is frequently compared to that of  $\text{Al}^{3+}$ , and there are several similarities that result from the fact that  $\text{Al}^{3+}$  and  $\text{Cr}^{3+}$  have charge to size ratios that are of similar magnitude.

Manganese forms three oxides,  $\text{MnO}$ ,  $\text{Mn}_2\text{O}_3$ , which is naturally occurring as the mineral *hausmanite*, and  $\text{MnO}_2$ , which is also the composition of the mineral *pyrolucite*. The lowest oxide,  $\text{MnO}$ , can be obtained by the partial reduction of  $\text{MnO}_2$  with hydrogen as shown in the following equation.

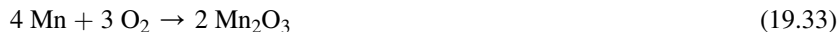


It can also be produced by the thermal decomposition of  $\text{Mn}(\text{OH})_2$ .

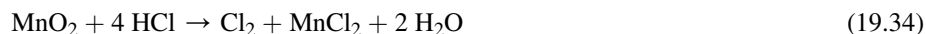


The structure of  $\text{MnO}$  consists of discrete  $\text{Mn}^{2+}$  and  $\text{O}^{2-}$  ions in a sodium chloride lattice.

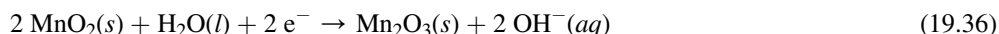
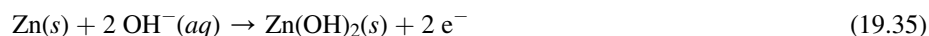
$\text{Mn}_2\text{O}_3$  can be prepared by the oxidation of the metal at high temperatures.



The most important of the manganese oxides is  $\text{MnO}_2$  that has the rutile structure (see Figure 4.6). This compound is the traditional oxidizing agent that is used in the laboratory preparation of chlorine by the reaction



$\text{MnO}_2$  has a number of uses in chemical processes as an oxidizing agent, and it is also used in dry cell and alkaline batteries. In both cases the anode is made of zinc. The anode reaction (oxidation) and cathode reaction (reduction) are as follows for an alkaline cell.



In this cell, a paste containing  $\text{KOH}$  is the electrolyte. In the older type of dry cell, the electrolyte is a moist paste of  $\text{NH}_4\text{Cl}/\text{ZnCl}_2$ . As has been mentioned in several places in this book, the  $\text{NH}_4^+$  ion is an acid so there is a slow reaction between the metal container and the electrolyte which causes the battery to deteriorate.

The oxide  $\text{Mn}_2\text{O}_7$  is a dangerous compound that is so unstable that it detonates, and it also reacts explosively with reducing agents such as organic compounds.

Iron oxides have been used for many centuries, not only as a source of the metal by reduction, but also as pigments. The relationship between iron oxides as products of the reaction of the metal with oxygen was discussed in Chapter 5. Only three iron oxides are significant enough to describe. The first of these is  $\text{FeO}$  that can be obtained by reactions that are frequently used to prepare metal oxides. Studies have shown that the compound is almost always deficient in iron to some extent. The fact that carbonates and oxalates decompose at elevated temperature to yield oxides has already been discussed.



The latter reaction is characteristic of many metal oxalates and can formally be written as

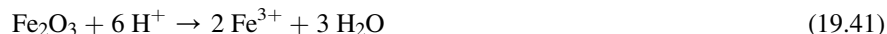


Under carefully controlled heating, oxalates can be decomposed to carbonates by loss of carbon monoxide.  $\text{FeO}$  also results when  $\text{Fe}$  reacts with  $\text{CO}_2$  at high temperature.



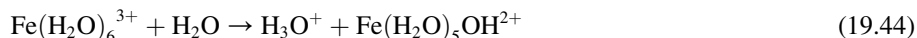
Magnetic iron oxide,  $\text{Fe}_3\text{O}_4$ , occurs naturally as the mineral *magnetite*. It has an inverse spinel structure (see Chapter 10) because it contains  $\text{Fe}^{2+}$  and  $\text{Fe}^{3+}$ , and the formula can be written as  $\text{FeO} \cdot \text{Fe}_2\text{O}_3$ .

The oxide containing  $\text{Fe}(\text{III})$  is  $\text{Fe}_2\text{O}_3$  but it also occurs as  $\text{Fe}_2\text{O}_3 \cdot \text{H}_2\text{O}$ , which has the same composition as  $\text{FeO}(\text{OH})$ . This oxide reacts with both acids and basic oxides as shown in the following equations:



Although they will not be discussed here, several forms of  $\text{Fe}_2\text{O}_3$  are known.

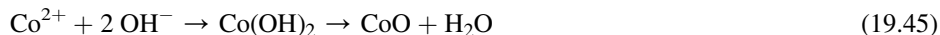
Solutions containing  $\text{Fe}^{3+}$  are acidic as a result of the reaction





This reaction is characteristic of hydrates of many metal ions having a small size and high charge because it reduces the charge density of the cation.

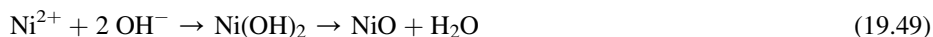
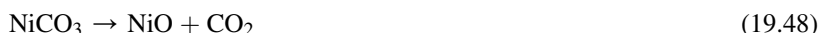
There are two known oxides of cobalt, CoO and Co<sub>3</sub>O<sub>4</sub>, but the existence of Co<sub>2</sub>O<sub>3</sub> is uncertain. The lowest of these, CoO, can be obtained by the decomposition of the hydroxide or carbonate containing Co<sup>2+</sup>.



In Co<sub>3</sub>O<sub>4</sub>, the Co<sup>2+</sup> ions are located in tetrahedral holes and the Co<sup>3+</sup> ions are located in octahedral holes of a spinel structure (see Section 10.2). This oxide can be prepared by the decomposition of Co(NO<sub>3</sub>)<sub>2</sub> because the oxide of cobalt that is more stable at high temperature is Co<sub>3</sub>O<sub>4</sub>.



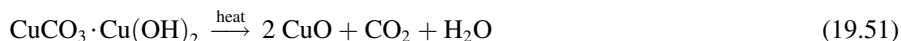
The only oxide of nickel that is very important is NiO, which can be prepared by the decomposition of the hydroxide or carbonate.



Copper forms two oxides, Cu<sub>2</sub>O and CuO, and of these the former is the more stable. It results when CuO is heated to a temperature above 1000 °C.



Cu<sub>2</sub>O also results when a basic solution containing Cu<sup>2+</sup> is reduced by a carbohydrate. In fact, this reaction is the basis of a test for sugars known as the *Fehling test*. This oxide also is used as an additive to glass because it imparts a red color. CuO can be obtained by heating the hydroxide or the carbonate in reactions that are analogous to those shown in Eqs (19.45) and (19.46). It also results from the decomposition of *malachite*, a beautiful green gemstone having the composition CuCO<sub>3</sub>·Cu(OH)<sub>2</sub>, when it is heated at moderate temperatures.



This oxide is used in making blue and green colored glass and in glazes for pottery.

Zinc forms only one oxide, ZnO, which is an amphoteric oxide, by reaction of the elements.



The reaction of ZnO with acids can be shown as



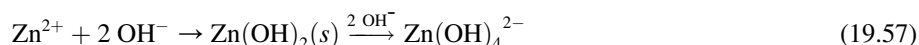
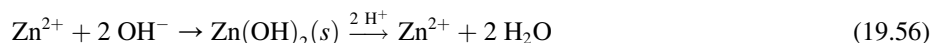
The reaction of ZnO with a base can be shown in different ways depending on whether the product is written as a *zincate* (ZnO<sub>2</sub><sup>2-</sup>) or a zinc hydroxide complex. If the reaction occurs between ZnO and a basic oxide such as Na<sub>2</sub>O, the equation can be shown as



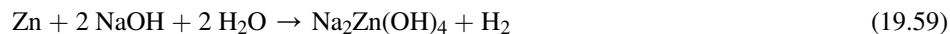
With an aqueous solution of a base, a zinc hydroxide complex forms.



If the metal hydroxide is precipitated first, the reactions with acid and base can be shown as



The metal will also dissolve readily in strong acids and strong bases.



A compound that has one color at room temperature and a different color at a high temperature is known as a *thermochromic* compound. Zinc oxide is such a compound because it is white at low temperature but yellow at high temperature. The structure of ZnO is like that of the *wurtzite* form of ZnS shown in Figure 4.5.

### 19.3 SULFIDES

Several ores containing transition metals, including copper, zinc, iron, and mercury among others, are found as the sulfides because of the low solubility of sulfide-containing compounds. Some minerals such as *chalcopyrite* ( $\text{CuFeS}_2$ ) are found as the sulfides containing at least two metals. Historically, many sulfide ores were used as pigments because of their bright, intense colors. Cadmium sulfide, CdS, is a bright yellow solid known as cadmium yellow, and a bright reddish-orange pigment known as vermilion is found as the mineral *cinnabar*, mercury(II) sulfide. The compound can be prepared by the reaction of mercury with melted sulfur. Initially, when HgS is prepared in this way, the color is black but, upon grinding, the colors of the pigments vary according to the size of the particles involved. It is a general characteristic of many compounds that the smaller the particle size, the lighter the color. Because of the toxicity of the heavy metals, many pigments containing these compounds have been replaced with other, less toxic pigments.

Some metal sulfides can be prepared by direct reaction of the elements or precipitation of an insoluble sulfide, but reaction of a metal sulfate with carbon may also be used.



Pure silver can react with hydrogen sulfide to form silver sulfide,  $\text{Ag}_2\text{S}$ , through a process known as tarnishing, which can be shown.



This process can also occur when silver items come in contact with sulfur-containing foods such as hard boiled eggs and mayonnaise.

For copper, there are two common sulfides,  $\text{Cu}_2\text{S}$  and  $\text{CuS}$ , but  $\text{Cu}_2\text{S}$  is more stable so when  $\text{CuS}$  is heated the following reaction occurs:



$\text{MoS}_2$  and  $\text{WS}_2$  possess three-dimensional layered structures in which the metal ion is surrounded by six sulfide ions in an octahedral arrangement and each sulfide ion is connected to three metal ions. The metal ion is sandwiched between layers of sulfide ions. Weak van der Waals attractions hold the layers together, and as a result,  $\text{MoS}_2$  can be used as a lubricant for high-temperature and high-pressure applications. These compounds also form nanotube and fullerene-like structures. The sulfides of niobium and tantalum form similar structures that have analogous properties.

### 19.4 HALIDES AND OXYHALIDES

Although halides of titanium in the +2, +3, and +4 states are known, only  $\text{TiCl}_4$  has important uses. High-temperature reactions by which  $\text{TiCl}_4$  can be produced include the following



One of the reasons for the importance of  $\text{TiCl}_4$  is that it reacts with  $[\text{Al}(\text{C}_2\text{H}_5)_3]_2$  to generate the effective catalyst that is used in the *Ziegler–Natta process* for the polymerization of ethylene (see Chapter 22). Halides such as  $\text{TiCl}_4$  hydrolyze in water as do most covalent halides.



Vanadium forms halides in the +2, +3, +4, and +5 oxidation states. For the +2 and +3 states, all of the halides are known, but only the fluoride is well characterized for the +5 state. Both  $VCl_4$  and  $VBr_4$  are rather rare compounds. Some of the vanadium halides can be prepared by the direct reaction of the elements as illustrated by the reaction



The halides of vanadium(IV) are unstable with respect to disproportionation so  $VF_4$  reacts to give  $VF_5$  and  $VF_3$ .

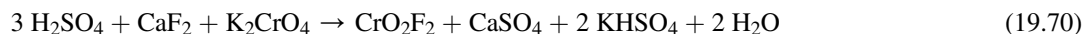


Transition metals in high oxidation states demonstrate some types of behavior resembling the chemical characteristics of nonmetals. For example, sulfur forms oxyhalides such as  $SOCl_2$  and  $SO_2Cl_2$  and vanadium in the +5 oxidation state behaves similarly in the formation of compounds having the formulas  $VOX_3$  and  $VO_2X$  (or  $OVX_3$  and  $XVO_2$  because both the oxygen and the halogen are bonded to vanadium).

Chromium halides having the formulas  $CrX_2$  and  $CrX_3$  (where  $X = F, Cl, Br, \text{ or } I$ ) are all known. The  $CrX_3$  compounds in particular can behave as Lewis acids toward a variety of electron pair donors. The following reactions can be used to prepare  $CrCl_3$ .



In analogy to  $SO_2Cl_2$  (sulfuryl chloride) and  $POCl_3$  (phosphoryl chloride), compounds having the formula  $CrO_2X_2$  are known as *chromyl* halides. The fluorides and chlorides are prepared by the reactions



If  $NaCl$  or  $KCl$  is substituted for  $CaF_2$  in the latter equation,  $CrO_2Cl_2$  is produced. The reactions of chromyl halides with water are typical of covalent halides, and they hydrolyze readily.



In addition to  $CrO_2F_2$ , an oxyfluoride having the formula  $CrOF_4$  is also known.

Although  $ReF_7$  and  $TcF_6$  are known, the highest oxidation state of Mn in a halide is +4 in the compound  $MnF_4$ . This observation shows the general trend that higher oxidation states are more commonly found for the heavier members of a group.  $MnO_3F$  and  $MnO_3Cl$  are known, but they are not important compounds. The tetrahalide can be prepared by the reaction of the elements as illustrated for the fluoride.



The trihalide,  $MnF_3$ , has also been characterized.

Iron generally forms only two series of halides,  $FeX_2$  and  $FeX_3$ , except that  $FeI_3$  is not stable due to the oxidizing nature of  $Fe^{3+}$  and the reducing character of  $I^-$ . The compound decomposes as illustrated by the reaction



The dichloride can be obtained as a hydrate as represented by the reaction



Evaporation of the solution yields  $FeCl_2 \cdot 4 H_2O$ , but, as in the case of many other aqueous halides, continued heating the solid leads to decomposition rather than dehydration. However, the anhydrous dichloride can be prepared by the reaction of iron with gaseous  $HCl$ .



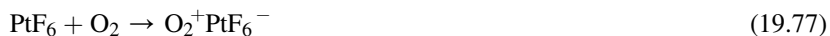
$FeCl_3$  is a Lewis acid and can function as a catalyst for many organic reactions. Trihalides of iron can be prepared by the reaction



The oxyhalides of iron are not of much importance.

Although  $\text{CoF}_3$  exists, the other halides of Co(III) are not stable because  $\text{Co}^{3+}$  is a strong oxidizing agent. Even the fluoride is so reactive that it is sometimes used as a fluorinating agent.

The only halides of nickel that are stable are those of nickel(II). Although we have not dealt specifically with the heavier metals in each group, this group deserves special mention. Palladium gives one halide in the +4 state,  $\text{PdF}_4$ , but platinum gives  $\text{PtF}_6$ . The latter is an extremely strong oxidizing agent that will oxidize  $\text{O}_2$  and Xe (see Chapters 15 and 18) as shown in the following equations:



Copper forms two series of halides, Cu(I) and Cu(II), but  $\text{CuI}_2$  is not stable because  $\text{Cu}^{2+}$  is an oxidizing agent and  $\text{I}^-$  is a reducing agent. As a result,  $\text{CuI}_2$  decomposes as illustrated by the equation



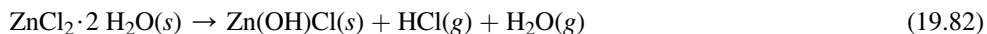
Because zinc has a  $3d^{10} 4s^2$  configuration, it routinely forms +2 compounds, and all of the halides having the formula  $\text{ZnX}_2$  are known. Of these,  $\text{ZnCl}_2$  is the most important because it is used in textile processing. The anhydrous compound is prepared by the reaction of zinc with gaseous HCl.



Although zinc dissolves readily in aqueous HCl according to the equation



evaporation of the solution yields  $\text{ZnCl}_2 \cdot 2 \text{H}_2\text{O}$ , and heating this solid results in the formation of  $\text{Zn(OH)Cl}(s)$  rather than anhydrous  $\text{ZnCl}_2$ .



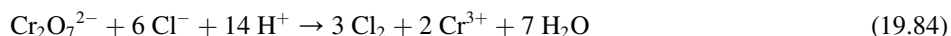
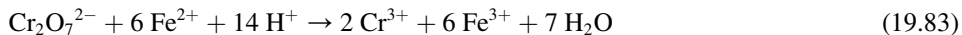
As indicated by this equation and others shown previously, it is frequently not possible to prepare an anhydrous metal halide by simply dehydrating a hydrate of the metal halide.

## 19.5 MISCELLANEOUS COMPOUNDS

There are some other compounds of the transition metals that are of sufficient importance to be mentioned in this brief survey. Consequently, in this section we will describe a few of the more important compounds of transition metals that do not fit in the categories discussed earlier.

Because  $\text{TiCl}_4$  is a strong Lewis acid, it forms many adducts with a wide variety of Lewis bases. As in the case of other covalent halides (e.g.,  $\text{PCl}_5$ ,  $\text{SO}_2\text{F}_2$ , etc.), it reacts with alcohols to form alkoxides,  $\text{Ti(OR)}_4$ . These compounds can be considered as “salts” of the acid  $\text{H}_4\text{TiO}_4$ .

Another compound that has wide utility as an oxidizing agent is potassium dichromate,  $\text{K}_2\text{Cr}_2\text{O}_7$ . This compound is an oxidizing agent that will oxidize a broad spectrum of materials in synthetic reactions, and it is used in redox titrations in analytical chemistry. The reactions with  $\text{Fe}^{2+}$  and  $\text{Cl}^-$  can be shown as follows.



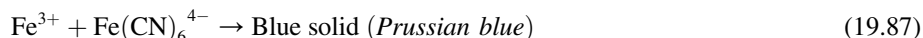
There are two oxyanions containing Cr(VI), the chromate ion,  $\text{CrO}_4^{2-}$  (which is yellow in color), and the dichromate ion,  $\text{Cr}_2\text{O}_7^{2-}$  (which is orange). The equilibrium between  $\text{CrO}_4^{2-}$  and  $\text{Cr}_2\text{O}_7^{2-}$  is pH dependent which results in a yellow color in basic solutions but the color changes to orange in acidic solutions. The equilibrium is represented by the equation



The chromates and dichromates are used in making pigments, in dyeing textiles, and in the tanning of leather. Chromium in the +6 oxidation state is believed to be a carcinogen so chromates and dichromates should be handled carefully.

Potassium permanganate,  $\text{KMnO}_4$ , is also a very useful oxidizing agent. Because of its deep purple color and the fact that the reduction product in acidic solutions,  $\text{Mn}^{2+}$ , is almost colorless, it serves as its own indicator in titrations. When  $\text{MnO}_4^-$  is used as an oxidizing agent in acidic solutions,  $\text{Mn}^{2+}$  is the reduction product, but in basic solutions  $\text{MnO}_2$  is produced.

Several iron compounds, some of which are  $\text{FeCO}_3$ ,  $\text{FeCl}_2$ , and  $\text{FeSO}_4$ , have industrial uses. The last of these is used in the manufacture of inks and dyes. Two interesting reactions involving iron compounds can be shown as follows.



Considerable controversy over the nature of the blue solids existed for many years. It now appears that the solids have the same composition,  $\text{Fe}_4[\text{Fe}(\text{CN})_6]_3 \cdot 16 \text{H}_2\text{O}$ , in which there are cyanide bridges between the  $\text{Fe}^{2+}$  and  $\text{Fe}^{3+}$  ions (see Chapter 20). This situation may be similar to the case of  $\text{Sb}_2\text{Cl}_8$  (or  $\text{SbCl}_3 \cdot \text{SbCl}_5$ ) in which there is rapid electron exchange that results in the antimony atoms being equivalent rather than existing as Sb(III) and Sb(V) (see Chapter 14).

Copper sulfate has many uses in dyeing processes, insecticides, and other industrial uses. The compound is obtained as the pentahydrate,  $\text{CuSO}_4 \cdot 5 \text{H}_2\text{O}$ . When this compound is heated, water is lost in stages so that the trihydrate and monohydrate are produced in separate stages before the anhydrous compound is obtained.

Zinc carbonate is found as the mineral *calamine*, and it is used in skin ointments as is zinc oxide. These are applied as creams to treat several types of skin rashes, and will be described briefly in Chapter 23.

A large number of ternary compounds of nickel and cobalt exist (including the carbonates, hydroxides, nitrates, sulfates, etc.), but the chemistry of these materials is typical of other compounds containing these anions.

## 19.6 THE LANTHANIDES

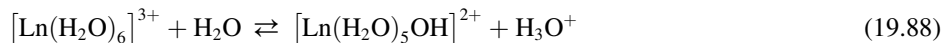
Although the lanthanides are not usually considered to be transition metals, they are sometimes referred to as “inner” transition metals because the third transition series begins with La and the next elements are the lanthanides. Accordingly, a brief introduction to their properties and chemistry is included here.

On the basis of orbitals filling as the sum  $(n + l)$  increases, following the  $6s$  orbital, for which  $(n + l) = 6$ , the next orbitals to fill would be expected to be those having the lowest value of  $n$  with  $(n + l) = 7$ . Consequently, those orbitals would be the  $4f$  set. The  $6s$  orbital is complete with barium ( $Z = 56$ ), but lanthanum ( $Z = 57$ ) has the electron configuration (Xe)  $5d^1 6s^2$  indicating that the  $5d$  orbitals start filling before the  $4f$ . Following La, cerium ( $Z = 58$ ) does not have the additional electron added to a  $d$  orbital to give the configuration  $5d^2 6s^2$  but rather the configuration is  $4f^2 6s^2$ . The  $4f$  level continues to fill until at europium ( $Z = 63$ ), the configuration is  $4f^7 6s^2$ . Gadolinium ( $Z = 64$ ) has the configuration  $4f^7 5d^1 6s^2$  as a result of the stability of the half-filled  $4f$  shell. Terbium ( $Z = 64$ ) has the configuration  $4f^9 6s^2$ , and the remainder of the elements in the series up to  $Z = 70$  have one additional electron in the  $4f$  level until it is complete at ytterbium. However, lutetium ( $Z = 71$ ) has the additional electron in the  $5d$  level to give  $4f^{14} 5d^1 6s^2$ . Therefore, except for relatively minor irregularities the lanthanides (sometimes indicated as a group by the general symbol Ln) fill the  $4f$  level in the elements cerium through lutetium. Several properties of the lanthanide elements are summarized in Table 19.4.

One of the interesting trends shown by the data in Table 19.4 concerns the gradual decrease in the sizes of the  $+3$  ions in progressing across the series. This effect is the result of the increase in nuclear charge and the fact that the  $4f$  levels are not effectively involved in screening. This decrease in size is sometimes referred to as the *lanthanide contraction*. Generally, the heats of hydration of cations increase as the charges on the ions increase and as the sizes decrease. The heat of hydration of the  $\text{Ln}^{3+}$  ions show this trend very well as illustrated in Figure 19.3.

Another interesting consequence of the lanthanide contraction is that the ionic radii of  $\text{Eu}^{3+}$  and  $\text{Ho}^{3+}$  are approximately the same as the radius of  $\text{Y}^{3+}$  (88 pm). As a result, much of the chemistry of  $\text{Y}^{3+}$  is similar to that of some of the lanthanides.

Cations that have a high charge to size ratio are strongly solvated in water, but they also undergo hydrolysis to relieve part of the high charge. Such reactions of the lanthanides can be represented by the equation



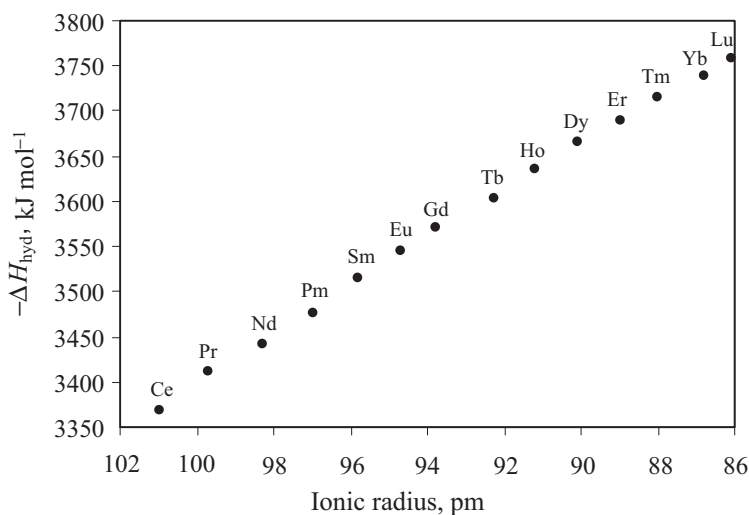
On the basis of the lanthanide contraction, it would be expected that solutions containing  $\text{Ln}^{3+}$  would increase in acidity across the group, and this prediction is correct.

Although some first-row transition metals are rather unreactive, the lanthanides are quite reactive and easily oxidized. This is illustrated by a comparison of the reduction potentials for lanthanum, magnesium, and sodium.



**TABLE 19.4** Properties of the Lanthanides

Metal	Structure <sup>a</sup>	m. p. °C	Ionic Radius (+3) <i>r</i> , pm	$-\Delta H_{\text{hyd}}$	First Three Ionization Potentials Sum, $\text{kJ mol}^{-1}$
Ce <sup>a</sup>	<i>fcc</i>	799	102	3370	3528
Pr	<i>fcc</i>	931	99.0	3413	3630
Nd	<i>hcp</i>	1021	98.3	3442	3692
Pm	<i>hcp</i>	1168	97.0	3478	3728
Sm	<i>rhmb</i>	1077	95.8	3515	3895
Eu	<i>bcc</i>	822	94.7	3547	4057
Gd	<i>hcp</i>	1313	93.8	3571	3766
Tb	<i>hcp</i>	1356	92.3	3605	3803
Dy <sup>a</sup>	<i>hcp</i>	1412	91.2	3637	3923
Ho	<i>hcp</i>	1474	90.1	3667	3934
Er	<i>hcp</i>	1529	89.0	3691	3939
Tm	<i>hcp</i>	1545	88.0	3717	4057
Yb <sup>a</sup>	<i>fcc</i>	824	86.8	3739	4186
Lu	<i>hcp</i>	1663	86.1	3760	3908

<sup>a</sup>Two or more forms are known.**FIGURE 19.3** Heat of hydration of +3 lanthanide ions as a function of ionic radius.

Some reduction potentials for other lanthanides are as follows: Ce,  $-2.48$  V; Sm,  $-2.40$  V; Ho,  $-2.32$  V; Er,  $-2.30$  V. Given the ease with which lanthanides are oxidized, many of their reactions with nonmetals are readily predictable. They react with halogens with the usual products being  $\text{LnX}_3$ . Reactions with oxygen (some of which are very vigorous!) lead to oxides having the formula  $\text{Ln}_2\text{O}_3$ . Like sodium and magnesium, some of the lanthanides are sufficiently reactive to replace hydrogen from water to produce the hydroxides,  $\text{Ln}(\text{OH})_3$  (as would be predicted by Eqs. (19.89)–(19.91)). The lanthanides have several important uses, and there are numerous interesting compounds of these elements. For coverage of these topics, the interested reader should consult the references listed.

In this chapter, we have discussed the importance of the transition metals from the standpoint of the metals themselves and a few of their compounds. Because the chemistry of the transition metals is so closely linked to their tendency and

ability to form coordination compounds, the next two chapters present an overview of that vast area of inorganic chemistry. Chapter 22 deals with the chemistry of organometallic compounds, many of which contain transition metals. Collectively, the last four chapters of this book show clearly the enormous importance of the transition metals and their chemical behavior.

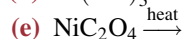
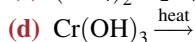
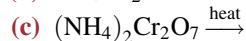
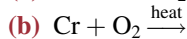
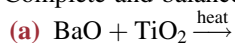
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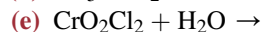
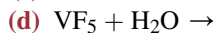
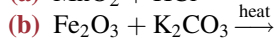
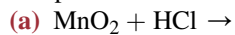
## PROBLEMS

- Predict which pairs of metals would be completely miscible.
  - Au/Ag
  - Al/Ca
  - Ni/Al
  - Ni/Co
  - Cu/Mg
- Explain why bismuth and cadmium are completely miscible as a liquid but almost completely insoluble as a solid.
- Two forms of CdS (wurtzite and zinc blende) exist, but CdO has the sodium chloride structure. What is one reason for this difference in behavior of these compounds?
- Explain why Cd(OH)<sub>2</sub> is a stronger base than Zn(OH)<sub>2</sub> which is amphoteric.
- Explain why there are many zinc compounds that are isostructural with those of magnesium.
- The apparent molecular weight of iron(III) chloride vapor is higher just above the boiling point (315 °C) than it is at 500 °C. Explain this observation.
- Write equations to show why heating CoCl<sub>2</sub>·6 H<sub>2</sub>O results in the formation of CoCl<sub>2</sub>(s) but heating CrCl<sub>3</sub>·6 H<sub>2</sub>O does not result in the formation of CrCl<sub>3</sub>.
- When a solution containing Zn<sup>2+</sup> has a solution of Na<sub>2</sub>CO<sub>3</sub> added, the precipitate is ZnCO<sub>3</sub>·Zn(OH)<sub>2</sub>. When a solution of NaHCO<sub>3</sub> is used instead, the precipitate is ZnCO<sub>3</sub>. Write the necessary equations and explain this difference.
- Explain why ZnO has the zinc blende structure but CdO has a structure of the sodium chloride type.
- Complete and balance the following.
  - $\text{ZnCO}_3 \cdot \text{Zn(OH)}_2(s) \xrightarrow{\text{heat}}$
  - $\text{Cr}_2\text{O}_7^{2-} + \text{Br}^- + \text{H}^+ \rightarrow$
  - $\text{ZnSO}_3(s) \xrightarrow{\text{heat}}$
  - $\text{MnO}_4^- + \text{Fe}^{2+} + \text{H}^+ \rightarrow$
  - $\text{ZnC}_2\text{O}_4 \xrightarrow{\text{heat}}$
- Given the facts that copper crystallizes in an *fcc* structure and the length of the unit cell is  $3.62 \times 10^{-8}$  cm, calculate the density of copper.
- Why is the solubility of copper in aluminum very low?

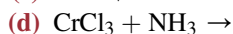
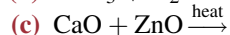
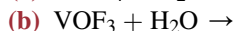
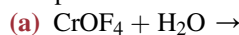
13. Complete and balance the following.



14. Complete and balance the following.



15. Complete and balance the following.



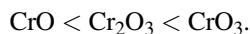
16. Show that the volume of free space in the *fcc* structure is 26%.

17. Draw structures for the chromate and dichromate ions.

18. Draw the structures for the ortho-, pyro-, and metavanadate ions.

19. Write equations to show the amphoteric behavior of chromium(III) hydroxide.

20. Explain why the acidic behavior of chromium oxides increases in the order



21. On the basis of atomic properties, explain why copper forms solid solutions that can have the following percent of lattice sites substituted for by the following atoms: Ni 100%, Al 17%, and Cr < 1%.

22. When 25% lead is dissolved in molten copper, a single liquid phase is present. If you were to examine the solidified mass, would you expect to find? Would this alloy be stronger or weaker than copper? Explain your answers.

23. On the basis of their properties, explain why separation of the lanthanides is possible but difficult.

24. On the basis of properties, explain why yttrium is frequently found with the lanthanide elements.

25. Write equations to show why aqueous solutions of  $\text{Ho}(\text{NO}_3)_3$  and  $\text{Nd}(\text{NO}_3)_3$  are acidic. Which would be more acidic, a 0.2 M solution of  $\text{Ho}(\text{NO}_3)_3$  or one of  $\text{Nd}(\text{NO}_3)_3$ ?

26. Using the data shown in Table 19.4, make a graph of the melting points of the lanthanides as a function of atomic number and explain any anomalies that appear in the graph.



# Structure and Bonding in Coordination Compounds

The chemistry of coordination compounds is a broad area of inorganic chemistry that has as its central theme the formation of coordinate bonds. A coordinate bond is the one in which both of the electrons used to form the bond come from one of the atoms, rather than each atom contributing an electron to the bonding pair, particularly between metal atoms or ions and electron pair donors. Electron pair donation and acceptance result in the formation of a *coordinate* bond according to the Lewis acid–base theory (see Chapter 6). However, compounds such as  $\text{H}_3\text{N}:\text{BCl}_3$  will not be considered as coordination compounds even though a coordinate bond is present between N and B. The term *molecular compound* or *adduct* is appropriately used to describe these “complexes” that are formed by interaction of molecular Lewis acids and bases. The generally accepted use of the term *coordination compound* or *coordination complex* refers to the assembly that results when a metal ion or atom accepts pairs of electrons from a certain number of molecules or ions. Such assemblies commonly involve a transition metal, but there is no reason to restrict the term in that way because nontransition metals ( $\text{Al}^{3+}$ ,  $\text{Be}^{2+}$ , etc.) also form coordination compounds.

There are numerous types of important coordination compounds (such as heme and chlorophyll) that occur in nature. Some coordination compounds are useful as a result of their ability to function as catalysts for industrially important processes. Also, the formation of coordination compounds is central to certain techniques in analytical chemistry. Accordingly, some understanding of the chemistry of coordination compounds is vital to students whose interests lie outside inorganic chemistry. Certainly, the field of coordination chemistry is much broader in its applicability than to just inorganic chemistry.

In Chapter 6, Lewis bases (electron pair donors) were classified as *nucleophiles* and electron pair acceptors were designated as *electrophiles* or Lewis acids. These concepts will now be used to describe coordination complexes of metals and numerous applications will be described in subsequent chapters.

## 20.1 TYPES OF LIGANDS AND COMPLEXES

When a metal ion is present in aqueous solutions, there is frequently a rather definite number of water molecules that are attached to the ion in a specific geometrical arrangement. For example, species such as  $\text{Fe}(\text{H}_2\text{O})_6^{3+}$ ,  $\text{Al}(\text{H}_2\text{O})_6^{3+}$ ,  $\text{Be}(\text{H}_2\text{O})_4^{2+}$ , etc., are considered to be the dominant species present in solutions that contain these metal ions. The bonds are the result of electron pairs being donated from the oxygen atoms in water molecules to empty orbitals on the metal ion. Such a situation is illustrated in Figure 20.1. The number of sites occupied by electron pair donors around the metal ion is known as the *coordination number*, and the groups attached to the metal are known as *ligands*. As will become apparent, there is a preference for coordination numbers of 2, 4, and 6, although some complexes exhibit coordination numbers of 3, 5, 7, and 8.

The water molecules attached directly to the metal ion form what is known as the *primary solvation* sphere. However, other water molecules in the vicinity of the hydrated ion are influenced by interactions with those molecules attached to the metal and thus have preferred orientations in order to maximize the hydrogen bonding forces between molecules. As a result, there is some structure of another layer of solvent molecules, which is called the *secondary solvation* sphere of the metal ion.

When the list of species that can potentially function as ligands is examined, it is a long one indeed. The number of molecules and ions that possess unshared pairs of electrons includes virtually all anions (such as  $\text{F}^-$ ,  $\text{Cl}^-$ ,  $\text{CN}^-$ ,  $\text{OH}^-$ ,  $\text{SCN}^-$ ,  $\text{NO}_2^-$ , etc.) and a large number neutral molecules (such as  $\text{H}_2\text{O}$ ,  $\text{NH}_3$ ,  $\text{CO}$ , amines, phosphines, sulfides, ethers, etc.). A few cations, e.g.,  $\text{H}_2\text{N}-\text{NH}_3^+$ , also have sites that can bind to metal ions by electron pair donation. Moreover, some molecules and ions have more than one atom that can function as an electron pair donor, which opens other possibilities. For example, the ethylenediamine molecule (usually abbreviated in formulas as en),  $\text{H}_2\text{NCH}_2\text{CH}_2\text{NH}_2$ , possesses

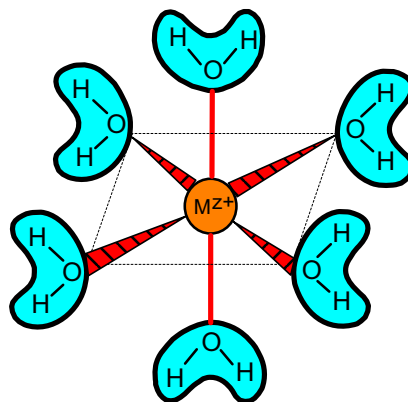
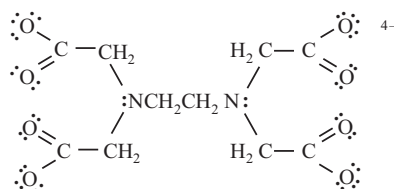


FIGURE 20.1 The hydration of a metal ion,  $M^{z+}$ , by six water molecules in the primary solvation sphere.

two nitrogen atoms that have unshared pairs of electrons. Both atoms can bind to the same metal ion to form a “ring” by occupying two adjacent bonding sites on the metal ion. Such a ring that contains the metal ion as one member is called a *chelate ring* (from the Greek word *chelos* meaning “claw”). A ligand that forms such a ring by attaching at two sites on the metal is called a *chelating agent*. A complex containing chelate rings is frequently referred to as a *chelate complex*.

Ethylenediamine has two electron donor atoms so it attaches at two points in the coordination sphere of the metal, and it is referred to as a *bidentate* ligand. The molecule known as triethylenediamine,  $H_2NCH_2CH_2NHCH_2CH_2NH_2$ , contains three electron pair donor atoms (the nitrogen atoms) so it normally attaches at three points in the coordination sphere of the metal, and it functions as a *tridentate* ligand. The anion of ethylenediaminetetraacetic acid (written as  $EDTA^{4-}$ ) has the structure

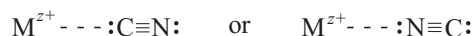


By utilizing four O atoms and two N atoms, the ethylenediaminetetraacetate ion can bind to six sites around a metal ion and is, therefore, a *hexadentate* ligand.

When the structure of  $CN^-$  is shown as



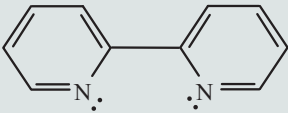
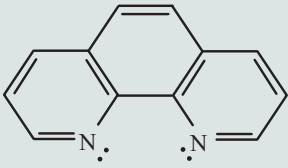
it can be seen that each end of the ion has an unshared pair of electrons so it is possible for the cyanide ion to bond to a metal ion through either end.



However, when the formal charges on the atoms are determined, it can be seen that the carbon end of the structure has a negative formal charge, and as a result it is usually that end of  $CN^-$  that bonds to metal ions. However, in some cases, the cyanide ion forms a bridge between two metal ions, especially in solid complexes, so that both ends are bonded to metal ions simultaneously. A similar situation exists for the thiocyanate ion,  $SCN^-$  because the ion can bond either through the sulfur atom or through the nitrogen atom. Compounds containing both types of linkages are known. The nitrite ion,  $NO_2^-$ , can bond through either the nitrogen atom or through an oxygen atom, and both types of complexes are known for this ligand. Ligands that can bond to a metal through more than one electron donor atom are said to be *ambidentate* ligands.

As has been mentioned, a very large number of molecules and ions are potential ligands. Table 20.1 lists a few of the common molecules and ions that form frequently encountered complexes. Because some ligands take on names in complexes that are different from their usual names (see Section 20.2), the names for the species when they function as ligands in coordination compounds are also shown.

**TABLE 20.1** Some Common Ligands

Ligand	Formula	Ligand Name
Ammonia	NH <sub>3</sub>	Ammine
Water	H <sub>2</sub> O	Aqua
Chloride	Cl <sup>-</sup>	Chloro
Cyanide	CN <sup>-</sup>	Cyano
Hydroxide	OH <sup>-</sup>	Hydroxo
Carbonate	CO <sub>3</sub> <sup>2-</sup>	Carbonato
Thiocyanate	SCN <sup>-</sup>	Thiocyanato
Nitrite	NO <sub>2</sub> <sup>-</sup>	Nitrito or nitro
Oxalate	C <sub>2</sub> O <sub>4</sub> <sup>2-</sup>	Oxalato
Carbon monoxide	CO	Carbonyl
Ethylenediamine	NH <sub>2</sub> CH <sub>2</sub> CH <sub>2</sub> NH <sub>2</sub>	Ethylenediamine
Acetylacetonate	$  \begin{array}{c}  \text{:O:} \quad \quad \quad \text{:}\ddot{\text{O}}\text{:}^- \\  \parallel \quad \quad \quad   \\  \text{H}_3\text{C}-\text{C}-\text{C}=\text{C}-\text{CH}_3 \\    \\  \text{H}  \end{array}  $	Acetylacetonato
2,2'-Dipyridyl		2,2'-Dipyridyl
1,10-Phenanthroline		1,10-Phenanthroline

## 20.2 NAMING COORDINATION COMPOUNDS

As in naming the enormous number of organic compounds, systematic procedures are necessary for deriving the names for coordination compounds. In the late 1800s, when the work on coordination compounds was just beginning, complexes were sometimes assigned names that were derived from the names of the investigators who prepared them. Thus,  $\text{K}[\text{C}_2\text{H}_5\text{PtCl}_3]$  became known as *Zeise's salt*;  $\text{NH}_4[\text{Cr}(\text{NCS})_4(\text{NH}_3)_2]$  was known as *Reinecke's salt*;  $[\text{Pt}(\text{NH}_3)_4][\text{PtCl}_4]$  was known as *Magnus' green salt*; and another isomer was known as *Magnus' pink salt*. Although some of these names are still encountered, as the number of coordination compounds increased this way of naming them proved inadequate.

Another procedure that was used in the early days of coordination chemistry was to refer to the compound by its color. For example,  $[\text{Co}(\text{NH}_3)_6]\text{Cl}_3$ , an intensely yellow compound, was known as “luteo cobaltic chloride” because the prefix *luteo* indicates the yellow color. Similarly,  $[\text{Co}(\text{NH}_3)_5\text{Cl}]\text{Cl}$  was known as “purpureo cobaltic chloride” because of its purple color, but this is not a satisfactory system of nomenclature to deal with a very large number of compounds.

The modern procedures used for naming coordination compounds are based on a few simple rules that will now be listed and then used to derive names.

1. In naming a coordination compound, the cation is named first followed by the name of the anion. One or both of these ions may be complex ions.
2. To name a complex ion, ligands are named first in alphabetical order. A prefix indicating the number of ligands is not considered to be part of the name of the ligand when determining alphabetical order. For example trichloro (three chloride ions) is named in the order indicated by the name chloro rather than by *tri*.
  - (a) Any coordinated anions end in *o*. Thus,  $\text{Cl}^-$  is chloro;  $\text{CN}^-$  is cyano;  $\text{SCN}^-$  is thiocyanato, etc.
  - (b) Neutral ligands are normally named using the name of the molecule. For example,  $\text{H}_2\text{NCH}_2\text{CH}_2\text{NH}_2$  is ethylenediamine;  $\text{C}_5\text{H}_5\text{N}$  is pyridine, etc., but a few ligands have special names. For example,  $\text{H}_2\text{O}$  is aqua,  $\text{NH}_3$  is ammine, and  $\text{CO}$  is carbonyl. Table 20.1 shows the names for many of the common ligands.
  - (c) Any cations that are coordinated end in *ium*. Although such cases are not particularly common, one example is hydrazine,  $\text{NH}_2\text{NH}_2$ , which can have one nitrogen function as a proton acceptor and still coordinate to a metal through the other nitrogen atom. In that case, the coordinated  $\text{NH}_2\text{NH}_3^+$  ion is named as hydrazinium.
3. To indicate the number of ligands of each type, the prefixes *di*, *tri*, and so on are used for simple ligands such as  $\text{Cl}^-$ ,  $\text{CN}^-$ , etc. For molecules that contain one of these prefixes in the name of the ligand, e.g., ethylenediamine, the prefixes *bis*, *tris*, *tetrakis*, and so on are used to indicate the number of ligands.
4. After naming the ligands, the name of the metal follows with its oxidation state indicated by Roman numerals.
5. If the complex ion is an anion, the name of the metal ends in *ate*.

Although other rules are needed for naming certain types of complexes, these rules are adequate for most cases. The use of these rules will now be illustrated by considering a few examples.

#### Example 1

Consider the compound  $[\text{Co}(\text{NH}_3)_6]\text{Cl}_3$ . The cation is  $[\text{Co}(\text{NH}_3)_6]^{3+}$  so it is named first. The ligands are ammonia molecules (named as ammine) and there are six of them. Therefore, the name is hexaamminecobalt(III) chloride.

#### Example 2

The compound  $[\text{Co}(\text{NH}_3)_5\text{Cl}]\text{Cl}_2$  contains five ammine ligands and one chloro ligand. Thus, the name for this compound is pentaamminechlorocobalt(III) chloride. Note that the prefix *penta* is not considered when determining the order in which the ligands are named.

#### Example 3

The compound  $\text{K}_4[\text{Fe}(\text{CN})_6]$  has been known for many years. The  $\text{CN}^-$  ligand is named as *cyano* so using the rules listed above, it can be seen that the name for  $\text{K}_4[\text{Fe}(\text{CN})_6]$  is potassium hexacyanoferrate(III).

#### Example 4

Many years ago, the compound  $\text{NH}_4[\text{Cr}(\text{CN})_4(\text{NH}_3)_2]$  was produced by a chemist named Reinecke. Although it has been called Reinecke's salt since that time, it has the systematic name that is derived using the rules stated earlier. Coordinated  $\text{NH}_3$  is named as *ammine* and  $\text{CN}^-$  is *cyano* so the correct name for Reinecke's salt is ammonium diamminetetracyanatochromate(III).

#### Example 5

A chemist named Magnus prepared two compounds that carried his name. A green salt,  $[\text{Pt}(\text{NH}_3)_4][\text{PtCl}_4]$ , was known as Magnus' green salt, and it contains both a complex cation and a complex anion. The cation contains a  $\text{Pt}^{2+}$  ion with four  $\text{NH}_3$  molecules attached so it has the name tetraammineplatinum(II). The anion has four  $\text{Cl}^-$  ions attached to a  $\text{Pt}^{2+}$  ion. Because the metal ion in this case is in an *anion*, its name ends in *ate*. Therefore, the name of the anion is tetrachloroplatinate(II), and the complete of the compound is tetraammineplatinum(II) tetrachloroplatinate(II).

A few other cases are listed below to give additional illustrations of applying the rules shown above.

$[\text{Co}(\text{NH}_3)_5\text{CN}](\text{NO}_3)_2$	Pentaamminecyanocobalt(III) nitrate
$[\text{Co}(\text{en})_2(\text{Br})_2]\text{Cl}$	Dibromobis(ethylenediamine)cobalt(III) chloride
$\text{K}[\text{Pt}(\text{C}_2\text{H}_4)\text{Cl}_3]$	Potassium trichloroethyleneplatinatate(II)

An additional rule, rule 6 below, is needed to deal with ligands that can bind to metals in more than one way.

6. Compounds in which there might be confusion as to how the ligand is attached have the bonding mode indicated by a letter after the name of the ligand. For example, it is possible for thiocyanate,  $\text{SCN}^-$ , to link to metal ions through either the nitrogen atom or the sulfur atom. The two bonding modes are distinguished by thiocyanato-N- and thiocyanato-S- in the name. Sometimes the mode of attachment is indicated by preceding the name with the letter (e.g., N-thiocyanato).

Some ligands contain two atoms that can function as electron pair donors so that the ligands can bind to two metal ions simultaneously forming a bridge. Rule 7 is needed to deal with such cases.

7. A bridging ligand is indicated by including  $\mu$  before the name of the ligand. Thus,  $[(\text{NH}_3)_3\text{Pt}(\text{SCN})\text{Pt}(\text{NH}_3)_3]\text{Cl}_3$  is named as hexammine- $\mu$ -thiocyanatodiplatinum(II) chloride.

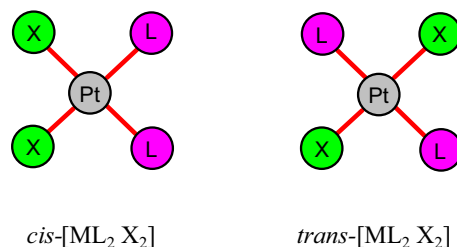
The simplified rules given above are sufficient to name the vast majority of coordination compounds, but other rules are needed in special cases. For example, organic compounds containing double bonds can frequently coordinate to metals in more than one way so the bonding mode must be incorporated in the name of the complexes. This type of complex will be described in Section 21.5.

## 20.3 ISOMERISM

Compounds having the same numbers and types of atoms but different structures are called *isomers*. Coordination compounds exhibit several of types of isomerism, and the study of these various types of isomers constitutes one of the interesting and active areas of research in coordination chemistry. Because so much of coordination chemistry is concerned with isomeric compounds, it is essential that a clear understanding of the various types of isomerism be achieved before a detailed study of structure and bonding in complexes is undertaken. Although the possibility of a substantial number of types of isomerism exists, only the more important types will be discussed here.

### 20.3.1 Geometrical Isomerism

Geometrical isomerism occurs when compounds have the same composition but different geometrical arrangements of atoms. One way this situation arises is when different groups can be positioned *cis* ( $90^\circ$ ) or *trans* ( $180^\circ$ ) to each other in a square planar complex. For a tetrahedral complex, all four bonding positions are equivalent so this possibility does not exist. As a result, a tetrahedral compound having the general formula  $\text{ML}_2\text{X}_2$  can have only one structure. However, if a compound  $\text{ML}_2\text{X}_2$  has a square planar arrangement, there are two possible structures, the *cis* and *trans* isomers.



In the case of an octahedral complex having the general formula  $\text{ML}_6$ , only one compound exists owing to the fact that all bonding positions in the coordination sphere of the metal are equivalent. For a complex having the formula  $\text{ML}_5\text{X}$ , there is still only one isomer possible for the same reason.

For a complex that has the formula  $\text{ML}_4\text{X}_2$  (as in the case of  $[\text{Co}(\text{NH}_3)_4\text{Cl}_2]^+$ ) there are possible structures other than octahedral. Compounds do not normally have a completely random arrangement of atoms but that still leaves the three possibilities shown in Figure 20.2.

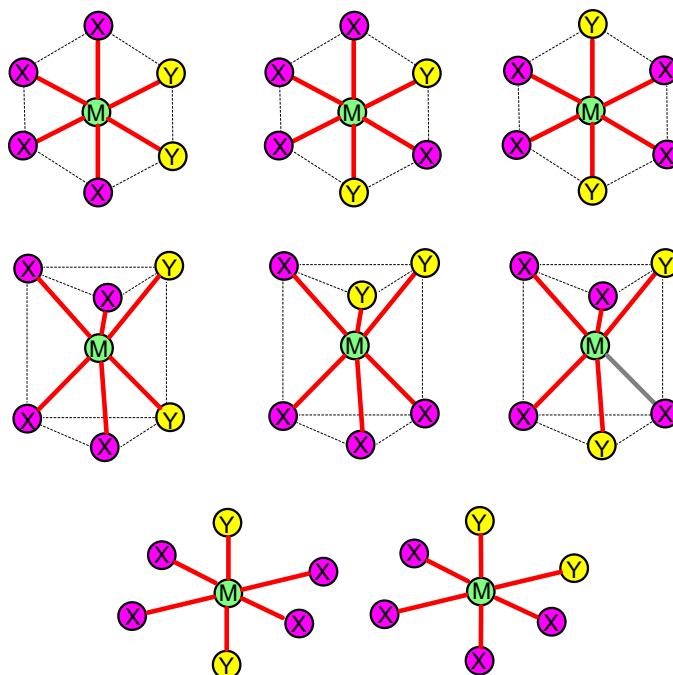
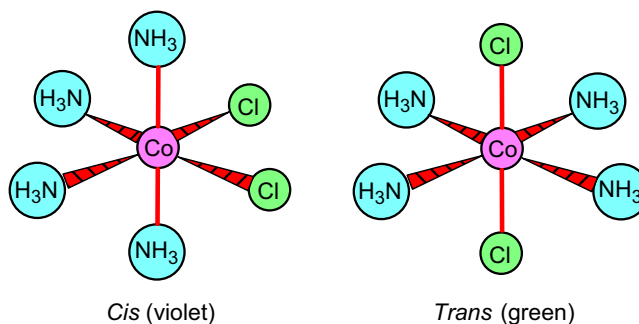


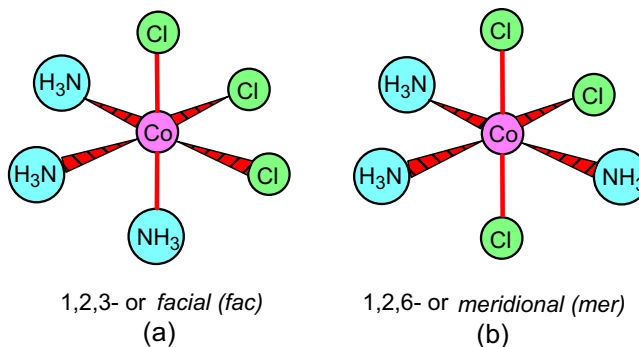
FIGURE 20.2 Isomers possible for a compound having the formula  $\text{MX}_4\text{Y}_2$  assuming planar hexagonal, trigonal prism, and octahedral structures.

When a complex having the formula  $[\text{Co}(\text{NH}_3)_4\text{Cl}_2]\text{Cl}$  was synthesized by Alfred Werner in the late 1800s, he found only two isomers, which led him to conclude that the structure was octahedral because either a planar hexagonal or trigonal prism structure would have resulted in three possible isomers. However, the octahedral structure could not be confirmed with the experimental techniques available at that time. More recently, X-ray analysis and spectroscopic methods have shown that the complex is actually octahedral.

For an octahedral complex having the general formula  $\text{ML}_4\text{X}_2$  the two possible isomers have *cis* and *trans* structures.



If the complex has the formula  $\text{ML}_3\text{X}_3$ , there are two possible isomers. For  $[\text{Co}(\text{NH}_3)_3\text{Cl}_3]$ , the two isomers have structures shown as follows.



In (a), the three chloride ions are on one face of the octahedron. In (b), the three chloride ions are occupying positions around an edge (a meridian) of the octahedron. Therefore, the names include *fac* and *mer* to indicate the structures.

As the number of different groups in the formula increases, the number of possible isomers increases rapidly. For example, if a complex has the general formula MABCDEF (where M represents a metal and A, B, ..., represent different ligands), a large number of isomers are possible.

### 20.3.2 Optical Isomerism

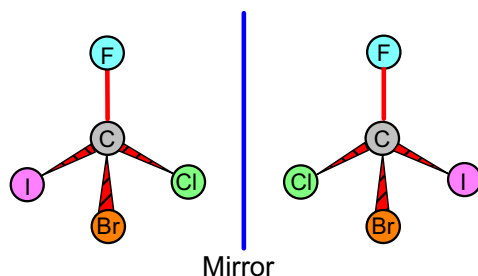
Optical isomers occur when a molecule and its mirror image is not superimposable. A simple molecule that has these properties is CFCIBrI. As shown in Figure 20.3, the two forms are illustrated on either side of a mirror. When the structure on the right is rotated by 180° in order to make the iodine atoms have the same position, the positions of the Cl and Br atoms are reversed. Thus, the two structures are not superimposable.

It can be shown that for molecules that do not possess a plane of symmetry, the mirror images are not superimposable. It is a property of such molecules that they rotate a beam of polarized light. If the beam is rotated to the right (when looking along the beam in the direction of propagation) the substance is said to be *dextrorotatory* (or simply *dextro*). Those substances that rotate the plane of polarized light to the left are said to be *levorotatory* (or simply *levo*). A mixture of equal amounts of the *dextro* and *levo* forms is called a *racemic mixture*, and it gives no net rotation of polarized light. Isomers of complexes that rotate polarized light in opposite directions are said to exhibit *optical isomerism*.

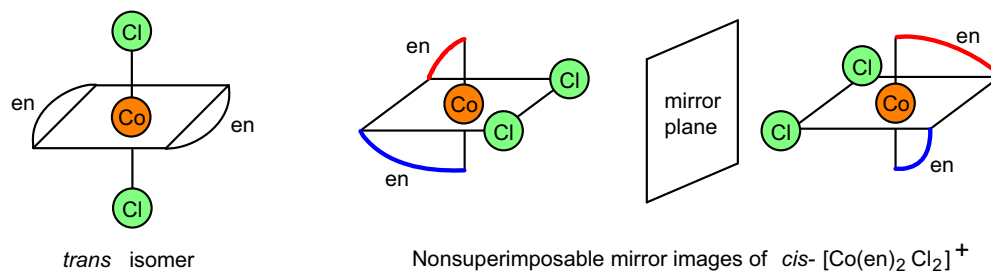
For the dichlorobis(ethylenediamine)cobalt(III) ion,  $[\text{Co}(\text{en})_2\text{Cl}_2]^+$ , two geometrical isomers, *cis* and *trans*, are possible. As was described in Chapter 3, a plane of symmetry (mirror plane) divides a molecule into equal fragments. For the *trans* isomer (shown in Figure 20.1) there is a plane of symmetry that bisects the cobalt ion and the ethylenediamine ligands with one chloride ion on either side. As shown in Figure 20.4, the *cis* isomer has no plane of symmetry and therefore, it can exist in *dextro* and *levo* forms.

### 20.3.3 Linkage Isomerism

Although the existence of *linkage isomers* is usually illustrated by recourse to the original example of the cobalt complexes  $[\text{Co}(\text{NH}_3)_5\text{NO}_2]^{2+}$  and  $[\text{Co}(\text{NH}_3)_5\text{ONO}]^{2+}$  studied by S. M. Jørgensen, a great many other cases are now known. Several



**FIGURE 20.3** The two optical isomers of CFCIBrI. Rotating the structure on the right by 180° so the Cl atoms could be superimposed, results in a structure in which the I and Br atoms are interchanged so the structures are not superimposable.

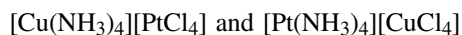
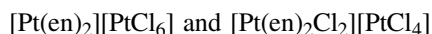
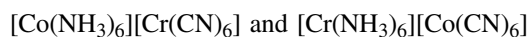


**FIGURE 20.4** The three isomers of  $[\text{Co}(\text{en})_2\text{Cl}_2]^+$ . If the structure on the right-hand side of the mirror is rotated by 180° to make the Cl atoms coincide, the rings have the opposite orientations to those shown in the structure on the left-hand side of the mirror.



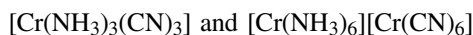
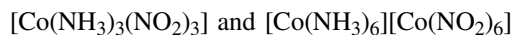


complexes having the same composition are  $[\text{Co}(\text{NH}_3)_5\text{CN}][\text{Co}(\text{NH}_3)(\text{CN})_5]$  and  $[\text{Co}(\text{NH}_3)_4(\text{CN})_2][\text{Co}(\text{NH}_3)_2(\text{CN})_4]$ , both of which still provide a coordination number of six for each metal. Other examples of this type of isomerism include the pairs shown below.



### 20.3.6 Polymerization Isomerism

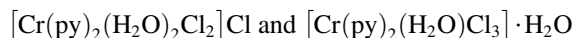
This type of isomerism is actually named incorrectly because polymers are usually not involved at all. The name stems from the fact that the composition of a polymer (an aggregate of monomer units) is the same as that of a monomeric unit that has a lower formula weight. In complexes, the term *polymerization isomerism* refers to the fact that a larger formula unit has the same overall composition as a smaller unit. Thus,  $[\text{Pd}(\text{NH}_3)_4][\text{PdCl}_4]$  has the same empirical formula as  $[\text{Pd}(\text{NH}_3)_2\text{Cl}_2]$ . With regard only to the empirical formula,  $[\text{Pd}(\text{NH}_3)_4][\text{PdCl}_4]$  is a “polymer” of the compound  $[\text{Pd}(\text{NH}_3)_2\text{Cl}_2]$  that has one-half the formula weight but the same composition. Other examples of this type of isomerism are illustrated by the following pairs.



The “polymer” is in reality only a complex having a higher formula weight and the same composition as a simpler one, but it does not involve repeating units as in the case of polymeric materials.

### 20.3.7 Hydrate Isomerism

Because many complexes are prepared in aqueous solutions, they are frequently obtained as crystalline solids that contain water of hydration. Water is, of course, also a potential ligand. As a result, isomeric compounds can sometimes be obtained in which water is coordinated in one case but is present as water of hydration in another. Compounds that differ in this way are called *hydrate isomers*. Because water is a neutral molecule, there must be one anion that is also held in a different way in the isomers. An example of this type of isomerism is illustrated by the compounds  $[\text{Cr}(\text{H}_2\text{O})_6]\text{Cl}_3$  and  $[\text{Cr}(\text{H}_2\text{O})_5\text{Cl}]\text{Cl}_2 \cdot \text{H}_2\text{O}$ . Other examples include the following pairs.



Although other types of isomerism in coordination compounds exist, the types described in this section represent the most important types.

## 20.4 FACTORS AFFECTING THE STABILITY OF COMPLEXES

### 20.4.1 The Nature of the Acid–Base Interaction

It has been recognized for many years that in a general way the basicity of the ligands has a great influence on the stability of complexes. After all, the formation of the coordinate bond is an acid–base reaction in the Lewis sense. However, as usually measured, basicity is toward the proton in aqueous solution. It sometimes provides a measure of the availability of electrons that might be expected when the ligands form coordinate bonds to metal ions.

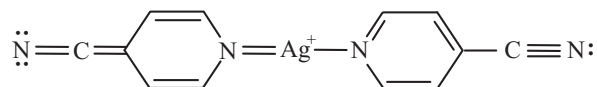
The basicity of a base, B, toward  $\text{H}^+$  is measured by the equilibrium constant,  $K_{\text{HB}}$ , for the reaction



In an analogous way, the coordination tendency of ligands toward silver ions is measured by the equilibrium constants for the following reactions.



In general, the relationship between the  $K$  values for the bases (measured toward the proton) is approximately linearly related to the  $K_1$  values for the formation of the silver complexes when several nitrogen bases (amines) having similar structure are considered. However, for a base such as 4-cyanopyridine, it is found that the silver complex is much more stable than would be predicted from the basicity of the ligand toward  $\text{H}^+$ . The reason for this is that some multiple bonding in the silver complex is possible as shown below.



However, such multiple bonding to  $\text{H}^+$  does not occur when it is attached to the ring nitrogen atom. Contributions from the structures having multiple bonds cause the ligand to bond more strongly to  $\text{Ag}^+$  than the basicity toward  $\text{H}^+$  would suggest. Another way to interpret the differences is in terms of the hard–soft interaction principle because  $\text{H}^+$  is a hard Lewis acid but  $\text{Ag}^+$  is a soft Lewis acid. Thus, one would expect a soft base (such as 4-cyanopyridine) to bond more strongly to  $\text{Ag}^+$  than would an amine (hard base) that bonds better to  $\text{H}^+$ .

In Chapter 6, the hard–soft interaction principle and its relationship to coordinate bonds was discussed. This principle can also be employed to predict which atom of an ambidentate ligand will coordinate to a given metal ion. For example,  $\text{SCN}^-$  usually coordinates to first row transition metal ions through the nitrogen atom, but toward second and third row metals the sulfur atom is usually the preferred bonding site. However, the nature of other ligands attached to the metal ion frequently alters the hard–soft character of the metal ion so that reversals occur. For example,  $\text{Co}^{3+}$  is usually considered as a hard Lewis acid. Accordingly, thiocyanate usually coordinates through the nitrogen atom as in  $[\text{Co}(\text{NH}_3)_5\text{NCS}]^{2+}$ . However, if the five other ligands are soft bases (e.g.,  $\text{CN}^-$ ), the preferred bonding mode of  $\text{SCN}^-$  is through the sulfur atom. As a result, in the case of  $\text{K}_3[\text{Co}(\text{CN})_5\text{SCN}]$ , the stable isomer in the solid state has S-bonded thiocyanate, and it does not isomerize to the N-bonded form. This change in the hard–soft character of the metal ion is referred to as a *sympiotic effect* in which the overall nature of the metal is altered by the other groups attached.

In a general way, it can be predicted that favorable bonding between a metal ion and ligands will occur when they have similar sizes and polarizabilities. Metal ions such as  $\text{Co}^{3+}$  and  $\text{Cr}^{3+}$  are hard Lewis acids and  $\text{NH}_3$  is a hard Lewis base, so it is expected that species such as  $[\text{Co}(\text{NH}_3)_6]^{3+}$  or  $[\text{Cr}(\text{NH}_3)_6]^{3+}$  would be stable. Uncharged metals are soft Lewis acids and CO is a soft Lewis base. Consequently, matching the hard–soft properties of the metal and ligands allows us to predict that  $\text{Fe}(\text{CO})_5$  will be a stable complex. Conversely, complexes such as  $[\text{Co}(\text{CO})_6]^{3+}$  (CO is a soft ligand) or  $[\text{Fe}(\text{NH}_3)_6]$  ( $\text{Fe}^0$  is a soft Lewis acid) would not be expected to be stable. Both sets of predictions are in accord with experimental observations.

### 20.4.2 The Chelate Effect

It is generally true that ligands that form chelates bind to metal ions more strongly than do monodentate ligands, even when the same donor atom is involved. For example, the formation constants for ethylenediamine ( $\text{H}_2\text{NCH}_2\text{CH}_2\text{NH}_2$ , a chelating agent abbreviated en) complexes are considerably greater than those in which ammonia molecules form complexes with the same metals. This is frequently referred to as the *chelate effect*. In each case, the donor atom is nitrogen and the basicities of ammonia and en are about the same. The values shown in Table 20.2 correspond to the reactions



and



Thus,  $K_1K_2K_3K_4$  for complexing by four  $\text{NH}_3$  molecules must be compared to  $K_1K_2$  for complexing by two en molecules because two of the latter occupy four coordination sites.

**TABLE 20.2** Comparison of Formation Constants for Ammine and Ethylenediamine Complexes

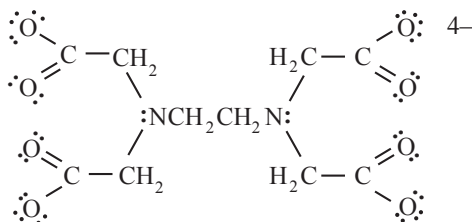
Metal Ion	Log( $K_1K_2K_3K_4$ ) $M(\text{NH}_3)_4^{2+}$	Log( $K_1K_2$ ) $M(\text{en})_2^{2+}$
Fe <sup>2+</sup>	3.70	7.72
Co <sup>2+</sup>	5.07	10.9
Ni <sup>2+</sup>	7.87	13.86
Cu <sup>2+</sup>	12.59	20.03
Zn <sup>2+</sup>	8.70	11.20

Thermodynamically, there is little difference between enthalpies of formation of bonds to NH<sub>3</sub> or en for a given metal. Both involve the formation of coordinate bonds between nitrogen atoms and the metal ions. However, the larger formation constants for the en complexes must reflect a more negative value of  $\Delta G$ . Recalling that

$$\Delta G = \Delta H - T\Delta S = -RT \ln K \quad (20.6)$$

it can be seen that a more negative  $\Delta G$  could arise from a more *negative*  $\Delta H$  or a more *positive*  $\Delta S$ . In this case, four NH<sub>3</sub> molecules coordinate to a metal ion in aqueous solution to replace four coordinated H<sub>2</sub>O molecules that become part of the bulk solvent. There is no change in the number of “free” or “bound” molecules. However, when en coordinates to a metal ion, *two* en molecules return *four* H<sub>2</sub>O molecules to the bulk solvent so that there is an increase in the number of free or unattached molecules. This increase in disorder results in a *positive*  $\Delta S$ , which in turn makes  $\Delta G$  more *negative*. Accordingly, the *chelate* effect is essentially a manifestation of increased entropy when en forms complexes compared to when NH<sub>3</sub> forms complexes.

Chelating agents that have more than two donor atoms can form additional rings so that one molecule of ligand can replace more than two molecules of solvent. For example, H<sub>2</sub>NCH<sub>2</sub>CH<sub>2</sub>NHCH<sub>2</sub>CH<sub>2</sub>NH<sub>2</sub> can coordinate through all three nitrogen atoms to form very stable complexes. Another very effective chelating agent is EDTA<sup>4-</sup>, the ethylenediamine-tetraacetate ion which has the structure



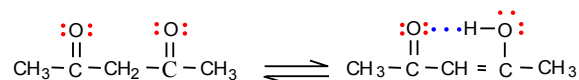
This ion forms complexes having very high stability, and it is widely used as an analytical reagent for analysis of metal ions by titration. It is also added in small amounts to salad dressings to complex with metal ions that are present in trace amounts. Such metal ions are known to catalyze the oxidation (spoilage) of the dressing. Entropy changes when such ligands replace several molecules of solvent are even more positive than when bidentate ligands are involved so the equilibrium constants for the replacement are very large.

### 20.4.3 Ring Size and Structure

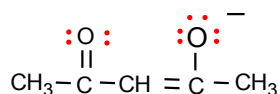
Although it has been shown that the chelate effect results in increased stability of metal complexes, there is ample evidence to show that some chelates are more stable than others having the same donor atoms. However, the stability of chelates is related to the number of atoms in the chelate rings. Chelate rings having five or six members usually result in the most stable complexes for a given series of ligands of the same type. For example, ethylenediamine forms chelates having five-membered rings (including, of course, the metal). A similar compound, 1,3-diaminopropane (pn) forms chelates also but the rings have six members and the complexes are generally less stable than those formed by en. Similarly, 1,4-diaminobutane shows even less tendency to form chelates. On the other hand, 1,2-diaminopropane forms chelates having stabilities to similar those of en complexes. It thus appears that five- or six-membered rings are the most stable, depending on the type of chelating agent.

In examining the complexes of a series of anions of dicarboxylic acids,  $^{-}\text{OOC}(\text{CH}_2)_n\text{COO}^{-}$ , it can be seen that the ring size depends on the number of  $\text{CH}_2$  groups. Thus, if  $n = 0$ , there are five atoms in the chelate ring; if  $n = 1$ , there are six; etc. Studies on the stability of complexes of this type show that 5- and 6-membered rings are of about equal stability, with a slight preference for the 5-membered rings (as is the case when  $n = 0$  for the oxalate ion,  $\text{C}_2\text{O}_4^{2-}$ ). There is, however, a rapid decrease in stability when  $n = 2, 3, \dots$ , for which 7, 8, ... membered rings result.

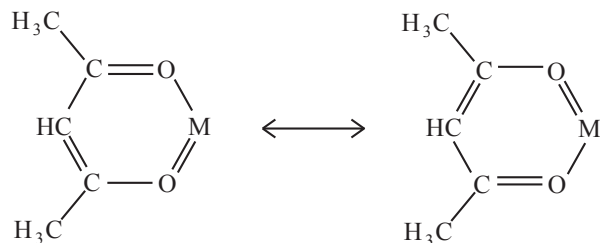
It is important to note that not only is ring *size* important, but also ring *structure*. For example, some of the most stable chelates are those of the acetylacetonate (acac) ion. Acetyl acetone (2,4-pentadione) undergoes a tautomerization reaction that can be shown as



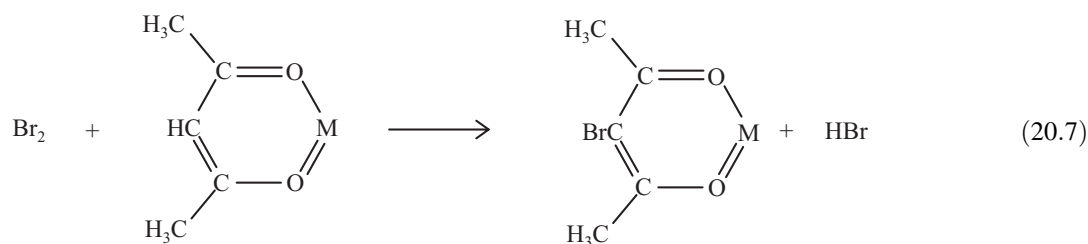
The enol form (on the right) is somewhat acidic and loses  $\text{H}^+$  easily to form the acetylacetonate anion (abbreviated as *acac*).



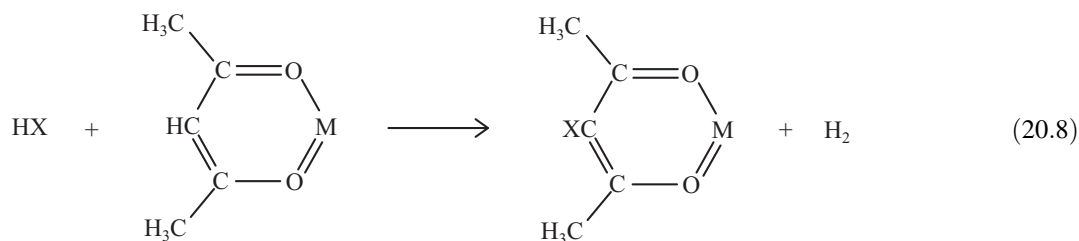
This ion complexes strongly with metal ions. When two such ions coordinate to a +2 metal ion, a neutral complex results and when three acac ions coordinate to a +3 metal ion, the complex is also neutral. The rings formed with the metal ion are planar with C–C and C–O bond distances of 138 and 128 pm, respectively. These bond distances are somewhat shorter than the values of 154 and 143 pm expected for single bonds of the C–C and C–O types. The planar chelate rings and the short bonds indicate that resonance structures can be drawn showing some  $\pi$ -electron delocalization. For simplicity, only one of the chelate rings in a complex such as  $\text{M}(\text{acac})_3$  is shown in the resonance structures below to display the bonding.



These structures illustrate that the rings possess some degree of aromaticity. Because acac complexes are usually uncharged and nonpolar, they are surprisingly soluble in organic solvents. The extreme stability of complexes of this type permits reactions to be carried out on the ring without disruption of the complex. One such reaction is bromination that is illustrated in Eqn (20.7).



Several reactions of this type, which are typically considered to be electrophilic substitutions, can take place on the chelate rings. The following equation can be considered as representing a general reaction of this type.



In this case,  $M = \text{Cr}^{3+}$ ,  $\text{Co}^{3+}$ , or  $\text{Rh}^{3+}$  and  $X = \text{CH}_3\text{CO}$ ,  $\text{HCO}$ , or  $\text{NO}_2$ . These observations indicate that both ring size and ring structure are important in determining the stability and properties of complexes.

## 20.5 A VALENCE BOND APPROACH TO BONDING IN COMPLEXES

Before beginning the discussion of bonding in complexes, their magnetic properties will be considered because the magnetic character of a complex tells us about how the electrons are distributed in the partially filled  $d$  orbitals. The essential measurement to determine the magnetism of a complex involves determining the attraction of the sample to a magnetic field. In one type of measurement, the sample is weighed in a special balance known as a Guoy balance in which the sample is suspended between the poles of a magnet. The sample is weighed with the magnetic field on and with it off. If the sample contains ions that have unpaired electrons, the electrons generate a magnetic field that is attracted to the applied magnetic field, and the sample is paramagnetic. Materials having all paired electrons have induced in them a magnetization that acts in opposition to the applied field. Therefore, if the sample has no unpaired electrons, it will be weakly repelled from the magnetic field and it is diamagnetic. Such a diamagnetic contribution is small compared to the attraction caused by paramagnetic behavior. Subtracting the diamagnetic contribution gives the *molar susceptibility*,  $\chi_M$ . A quantity that is related to the number of unpaired electrons is the *magnetic moment*,  $\mu$ . The magnetic moment is related to the molar magnetic susceptibility by the equation

$$\mu = \frac{3k}{N_0} \chi_M T \quad (20.9)$$

where  $k$  is Boltzmann's constant,  $N_0$  is Avogadro's number, and  $T$  is the absolute temperature.

For a single unpaired electron, the magnetic moment due to its spin is  $\mu_s = 1.73$  Bohr magneton (BM). The orbital motion may also contribute to the magnetic moment, and for a single electron the total magnetic moment is represented in terms of both the spin and orbital contributions by the equation

$$\mu_{s+l} = \sqrt{4s(s+1) + l(l+1)} \quad (20.10)$$

where  $s$  is the spin of one electron and  $l$  is the angular momentum quantum number for the electron. For several unpaired electrons having parallel spins, Eqn (20.10) becomes

$$\mu_{s+L} = \sqrt{4S(S+1) + L(L+1)} \quad (20.11)$$

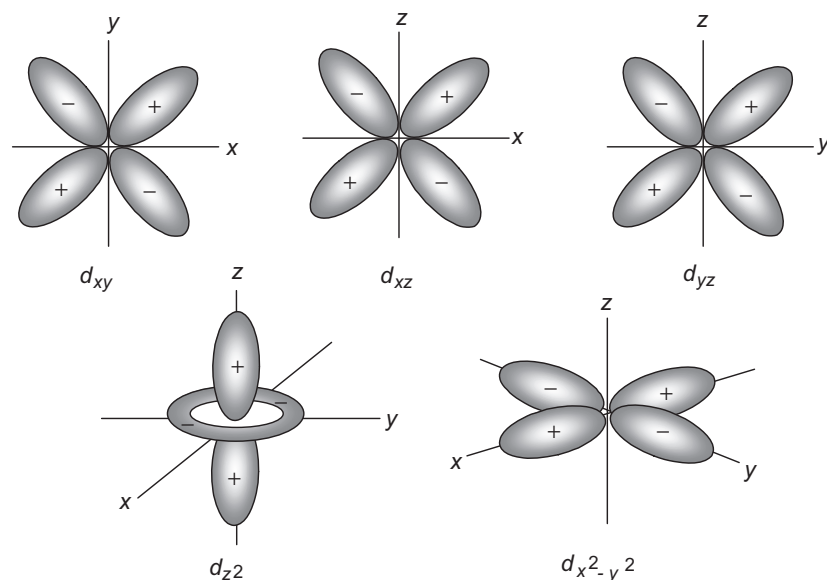
where  $S$  is the sum of the spins and  $L$  is the sum of the orbital angular momentum quantum numbers. Several factors may cause the orbital contribution to be much smaller than is predicted by Eqn (20.11). For many complexes of first row metal ions, the contribution is so small that it is ignored and the spin only magnetic moment is used. However, the sum of the spins,  $S$ , is related to the number of unpaired electrons,  $n$ , by  $S = n/2$ . Thus, the spin only magnetic moment can be written as

$$\mu_s = \sqrt{n(n+2)} \quad (20.12)$$

Table 20.3 shows the expected magnetic moments that correspond to ions having different numbers of unpaired electrons.

**TABLE 20.3** Spin Only Magnetic Moments

Number of Unpaired Electrons	S	$\mu_s$ (BM)
1	$\frac{1}{2}$	1.73
2	1	2.83
3	$\frac{3}{2}$	3.87
4	2	4.90
5	$\frac{5}{2}$	5.92

**FIGURE 20.5** A set of five  $d$  orbitals. The signs on the lobes correspond to the *mathematical* signs of the wave functions with reference to the coordinates.

Magnetic moments provide the most direct way to determine the number of unpaired electrons in a metal ion. When the number of unpaired electrons is known, it is possible to use that information to deduce the type of hybrid orbitals that are used by the metal ion in bonding.

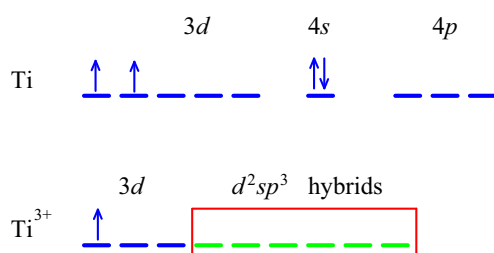
The  $d$  orbitals have orientations as shown in Figure 20.5, and atomic orbitals must have the appropriate orientations (symmetry) for hybrid orbitals to form. For example, there are no hybrid orbitals formed between the  $p_x$  and  $d_{yz}$  orbital (those orbitals are orthogonal). A few of the common types of hybrid orbitals involved in bonding in complexes are shown in Table 20.4.

The hybrid orbital type  $d^2sp^3$  refers to a case in which the  $d$  orbitals have a smaller principal quantum number than that of the  $s$  and  $p$  orbitals (e.g.,  $3d$  combined with  $4s$  and  $4p$  orbitals). The  $sp^3d^2$  hybrid orbital type indicates a case where the  $s$ ,  $p$ , and  $d$  orbitals all have the same principal quantum number (e.g.,  $4s$ ,  $4p$ , and  $4d$  orbitals) in accord with the natural order of filling atomic orbitals having a given principal quantum number. Some of the possible hybrid orbital combinations will now be illustrated for complexes of first row transition metals.

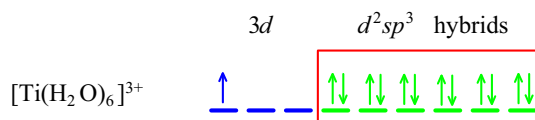
The titanium atom has the electron configuration  $3d^2 4s^2$ . Remembering that when atoms of transition elements lose electrons they are lost from the  $s$  orbital first, the  $Ti^{3+}$  ion has the configuration  $3d^1$ . The orbitals and electron populations for the Ti atom and  $Ti^{3+}$  ion can now be shown as below. After the loss of three electrons from Ti, two of the  $3d$  orbitals, the  $4s$ , and two of the  $4p$  orbitals form a set of six  $d^2sp^3$  hybrid orbitals.

TABLE 20.4 Important Hybrid Orbital Types in Complexes

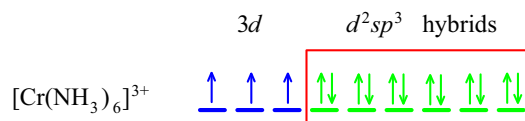
Atomic Orbitals	Hybrid Type	Number of Orbitals	Orientation
$s, p_x$	$sp$	2	Linear
$s, p_x, p_y$	$sp^2$	3	Trigonal planar
$s, p_x, p_y, p_z$	$sp^3$	4	Tetrahedral
$d_{z^2}, s, p_x, p_y$	$dsp^2$	4	Square planar
$d_{z^2}, s, p_x, p_y, p_z$	$dsp^3$	5	Trigonal bipyramid
$d_{x^2-y^2}, d_{z^2}, s, p_x, p_y, p_z$	$d^2sp^3$	6	Octahedral
$s, p_x, p_y, p_z, d_{x^2-y^2}, d_{z^2}$	$sp^3d^2$	6	Octahedral



The basic idea in this approach is that six *atomic* orbitals are used to form six *hybrid* orbitals that are directed toward the corners of an octahedron, the known structure of complexes containing  $Ti^{3+}$ . Hybrid orbital types that meet these requirements are  $d^2sp^3$  and  $sp^3d^2$ . In the case of the  $Ti^{3+}$  ion, two of the 3d orbitals are empty so the  $d_{x^2-y^2}$  and  $d_{z^2}$ , the 4s orbital, and the three 4p orbitals can form a set of six  $d^2sp^3$  hybrids. Moreover, the orbitals will be *empty* as is required for the metal to accept the six pairs of electrons from the six ligands. The electrons in the complex ion  $[Ti(H_2O)_6]^{3+}$  can now be represented as follows in which the six pairs of electrons from the ligands occupy  $d^2sp^3$  hybrid orbitals.



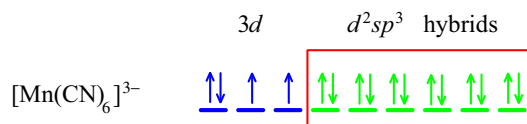
Other  $d^1$ ,  $d^2$ , or  $d^3$  metal ions would behave similarly because they would still have two empty 3d orbitals that could form a set of  $d^2sp^3$  hybrids. Thus, complexes of  $Ti^{2+}$ ,  $V^{3+}$ , and  $Cr^{3+}$  would be similar except for the number of electrons in the unhybridized 3d orbitals. The  $Cr^{3+}$  ion has a  $d^3$  configuration so three electrons will populate the 3d orbitals that are not used in hybridization. Addition of six ligands such as  $NH_3$  or  $H_2O$  will result in an octahedral complex in which the six pairs of donated electrons are arranged as shown below for  $Cr^{3+}$ , a  $d^3$  ion.



Chromium complexes such as  $[Cr(NH_3)_6]^{3+}$  will be paramagnetic and will have magnetic moments that correspond to three unpaired electrons per metal ion. The observed magnetic moments may not be exactly the predicted value of 3.87 BM, but they are sufficiently close so that three unpaired electrons are indicated.

For a complex containing a  $d^4$  ion there are two distinct possibilities. The four electrons may either reside in three available 3d orbitals and a set of  $d^2sp^3$  orbitals will be formed or else the empty 4d orbitals will be used to form a set of  $sp^3d^2$  hybrids. The first case would have two unpaired electrons per complex ion, whereas the second would have four unpaired electrons per complex ion. Consequently, the magnetic moment can be used to distinguish between these two cases. An example of a  $d^4$  ion is  $Mn^{3+}$ , for which two types of complexes are known. The first type of complex is

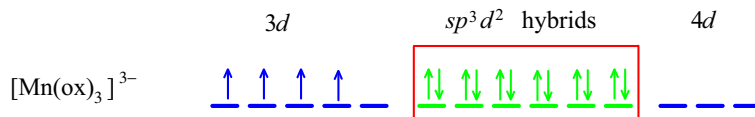
illustrated by  $[\text{Mn}(\text{CN})_6]^{3-}$ . The orbital population that results after the addition of six cyanide ions to form the complex  $[\text{Mn}(\text{CN})_6]^{3-}$  may be shown as follows.



This hybrid orbital type indicates that the complex would have two unpaired electrons that would result in a magnetic moment of about 2.83 BM. Although the actual magnetic moment of 3.18 BM is somewhat higher, it is sufficiently close to indicate that the complex contains two unpaired electrons.

If there were four unpaired electrons in the  $3d$  orbitals of  $\text{Mn}^{3+}$ , a magnetic moment of 4.90 BM would be predicted, and the  $d$  orbitals used in bonding would have to be  $4d$  orbitals giving a hybrid orbital type of  $sp^3d^2$ . Such a complex is referred to as an *outer orbital* complex because the  $d$  orbitals used are outside the normal valence shell of the metal ion. Because the electrons in the  $3d$  orbitals remain unpaired, the complex is also called a *high spin* complex. The  $[\text{Mn}(\text{CN})_6]^{3-}$  complex considered earlier is called a *low spin* or *inner orbital* complex. A ligand such as  $\text{CN}^-$  that forces the pairing of electrons is called a *strong field* ligand. Ligands such as  $\text{H}_2\text{O}$ ,  $\text{Cl}^-$ , or oxalate ( $\text{C}_2\text{O}_4^{2-}$  abbreviated as ox) that do not force electron pairing are called *weak field* ligands.

The electron arrangement in an *outer orbital* complex such as  $[\text{Mn}(\text{ox})_3]^{3-}$  (in which ox represents the oxalate ion,  $^- \text{OOC}-\text{COO}^-$ ) can be shown as follows.



The magnetic moment expected for a complex having four unpaired electrons is 4.90 BM, and the magnetic moment of this complex is about 4.81 BM, indicating that this complex is of the high spin or outer orbital type. As a result of the set of six hybrid orbital being composed of  $d$  orbital having the same principal quantum number (4) as the  $s$  and  $p$  orbitals, they are designated as  $sp^3d^2$  and a complex of this type is called an *outer orbital* complex.

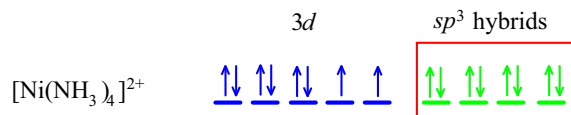
Metal ions having a  $d^5$  configuration have the same possibilities with regard to electron pairing as  $d^4$ . Accordingly, two types of complexes (high and low spin) are known for  $\text{Fe}^{3+}$ , a  $d^5$  ion, and they are easily distinguished by measuring the magnetic moment.

If only octahedral complexes are considered, it can be seen that two of the  $3d$  orbitals are required for the formation of  $d^2sp^3$  hybrid orbitals. The remaining three orbitals can hold a maximum of six electrons, so it is impossible to have two of the  $3d$  orbitals vacant if there are seven or more electrons in the metal ion. As a result,  $d^7$ ,  $d^8$ ,  $d^9$ , or  $d^{10}$  ions in octahedral complexes are always of the outer orbital (high spin) type.

For a  $d^6$  ion such as  $\text{Co}^{3+}$ , a low spin complex would contain no unpaired electrons, whereas the high spin case would contain four unpaired electrons. The low spin complex  $[\text{Co}(\text{NH}_3)_6]\text{Cl}_3$  has a magnetic moment of 0, whereas that of  $\text{K}_3[\text{CoF}_6]$  is 4.36 BM in accord with these predictions.

There is a strong tendency for  $d^7$ ,  $d^8$ ,  $d^9$ , and  $d^{10}$  ions to form complexes having structures that are not octahedral. One reason is that ions having these configurations frequently are +1 or +2 charged ( $\text{Co}^{2+}$ ,  $\text{Ni}^{2+}$ ,  $\text{Cu}^{2+}$ ,  $\text{Ag}^+$ , etc.). As a result, adding six pairs of electrons around the metal ion causes it to acquire considerable negative charge. Metal ions having a +3 charge have a greater charge to size ratio, which causes these ions to more readily attach to six ligands.

Consider the  $\text{Ni}^{2+}$  ion that has the configuration  $3d^8$ . It is readily apparent that if an octahedral complex is formed, it must be of the outer orbital type because two of the  $3d$  orbitals cannot be vacated to form  $d^2sp^3$  hybrids. However, by hybridization of the  $4s$  and  $4p$  orbitals a set of  $sp^3$  hybrids could be formed without involving the  $3d$  orbitals. It is not surprising to find that tetrahedral complexes of  $\text{Ni}^{2+}$  (such as  $[\text{Ni}(\text{NH}_3)_4]^{2+}$ ) are known that have magnetic moments of about 3 BM corresponding to two unpaired electrons. This type of complex involves hybrid orbitals that require no electron pairing and can be represented as







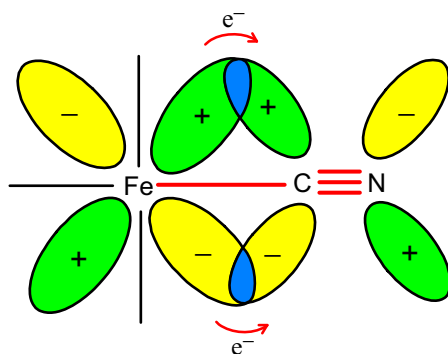


FIGURE 20.7 Back donation of electron density from  $\text{Fe}^{2+}$  to  $\text{CN}^-$  in a complex such as  $[\text{Fe}(\text{CN})_6]^{4-}$ .

where  $N_b$  is the number of electrons in bonding states and  $N_a$  is the number of electrons in antibonding states. However, there are  $\pi^*$  orbitals on  $\text{CN}^-$  that are empty. These orbitals have the directional properties shown in Figure 20.7.

In accepting six pairs of electrons donated by the ligands, the metal ion acquires a negative formal charge. Considering half of each bonding pair of electrons as belonging on the metal ion, a  $+3$  metal ion would acquire a  $-3$  formal charge when six ligands are attached. In order to eliminate part of this electron density from the metal, donation of electrons occurs from the  $d_{yx}$ ,  $d_{xz}$ , and  $d_{yz}$  metal orbitals to the empty  $\pi^*$  orbitals on the ligands. *Back donation* (sometimes called *back bonding*) occurs when the ligands possess vacant orbitals having suitable symmetry that are capable of accepting electron density from the metal. This results in some multiple bond character in the metal-ligand bonds.

By considering a complex such as  $\text{Fe}(\text{CN})_6^{3-}$ , it can be seen that the  $\text{Fe}^{3+}$  ion has accepted six pairs of electrons donated by the six  $\text{CN}^-$  ions. As a result, dividing the shared pairs equally to calculate the formal charge on the  $\text{Fe}^{3+}$  indicates that the formal charge is  $-3$  because the  $\text{Fe}^{3+}$  has “gained” six electrons and it has a  $+3$  ionic charge. As a result of gaining this much electron density, there is a tendency for the metal ion to try to relieve part of the negative charge. In this case, donation of some electron density back to the  $\text{CN}^-$  ions helps to reduce the negative charge on the metal ion. As shown in Figure 20.7, this can be accomplished because the symmetry of the  $\pi^*$  orbitals on  $\text{CN}^-$  matches that of the  $d$  orbitals of  $\text{Fe}^{3+}$ .

The  $d_{xy}$ ,  $d_{yz}$ , and  $d_{xz}$  orbitals are not used in forming  $d^2sp^3$  hybrid bonding orbitals, so electron density in them can be donated to the empty  $\pi^*$  orbitals on the cyanide ions. Because this electron donation is from the metal to the ligand, it is in the *reverse* direction to that which normally occurs when a coordinate bond forms. As a result, it is called *back donation*. This donation is not extensive, but it is significant enough that the  $\pi^*$  orbitals are populated to some degree. This electron population in the antibonding orbitals of  $\text{CN}^-$  reduces the bond order by increasing the value of  $N_a$  in Eqn (20.15). Reduction of the bond order in the  $\text{CN}^-$  ion causes the bond to be slightly longer and weaker than it is in a cyanide ion that is not coordinated to the metal ion. Accordingly, the position of the peak corresponding to stretching the C–N bond in the complex occurs at a lower frequency in the infrared spectrum than does that for ionic cyanide compounds. This type of spectral analysis gives us valuable information about the extent of this back donation. Of course, in order for back donation to occur, the metal ion must have electrons populating the nonbonding  $d$  orbitals and the ligand must have empty orbitals of appropriate symmetry to overlap with the nonbonding metal  $d$  orbitals. For a given number of coordinated ligands, back donation is less extensive when the metal ion has a higher positive charge because the metal will have a lower negative formal charge in this case.

The back donation of electron density from the metal to the  $\text{CN}^-$  results in a weakening of the C–N bond, but it gives additional strength to the Fe–C bond. In fact, the C–N bond is lengthened and weakened, but the Fe–C bond is shortened and strengthened. The partial back donation can be shown in terms of the resonance structures (with only one ligand being shown)



Although some double bonding between  $\text{Fe}^{3+}$  and  $\text{CN}^-$  occurs, the two structures do not contribute equally, and the actual structure more closely resembles the structure on the left. Also, the back donation is spread among all six  $\text{CN}^-$  ligands.

Cyanide is not the only ligand that can accept electron density from the metal in this way. For example, a ligand in which phosphorus is the donor atom can also accept back donation. In this case, it is the empty  $d$  orbitals on the P atom that accept the electrons. Because of the ability of  $\text{CN}^-$  to accept electron density, it behaves as a particularly good ligand, and it forces electron pairing in cases where a ligand that cannot interact with the metal ion in this way might not. Other types of ligands that lend themselves to this multiple bonding are CO, phosphines, ethylene, and other ligands having  $\pi$ -bonding abilities because of their ability to accept electron density from the nonbonding metal orbitals in empty ligand orbitals of appropriate symmetry.

## 20.7 LIGAND FIELD THEORY

The valence bond approach described in Section 20.5 is useful for explaining some properties of complexes, but it does not provide an explanation for many other observations. For example, why does  $[\text{Mn}(\text{H}_2\text{O})_6]^{3+}$  have four unpaired electrons but  $[\text{Mn}(\text{CN})_6]^{3-}$  has only two? According to the valence bond descriptions of these complexes,  $[\text{Mn}(\text{H}_2\text{O})_6]^{3+}$  has  $sp^3d^2$  hybrid orbitals, whereas  $[\text{Mn}(\text{CN})_6]^{3-}$  has  $d^2sp^3$ . What is the difference between  $\text{H}_2\text{O}$  and  $\text{CN}^-$  that causes pairing of electrons in one case but not in the other? So, the question remains as to why the bonding is different in these cases. Another description of the bonding in metal complexes is known as ligand field theory, and a simplified presentation of that approach will be presented.

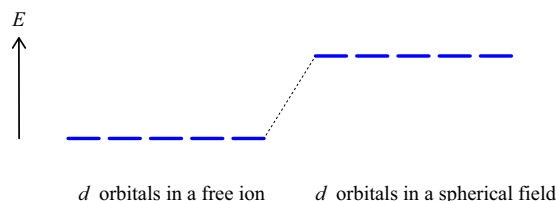
Crystal field theory was developed by a physicist, Hans Bethe, in 1929 while studying the spectral characteristics of metal ions in crystals. The negative ions generate an electrostatic field around the metal ions and crystal field theory strictly deals only with the electrostatic interactions. The model was adapted by J. H. Van Vleck to coordination compounds in 1930s because a metal ion in a complex is surrounded by anions or polar molecules in an environment that resembles that in which the metal ion is found in a crystal. Because the field in a complex is generated by ligands, the field is referred to as a *ligand field*. Consequently, ligand field theory presents a description of what happens to the valence orbitals of a positive transition metal ion when it is surrounded by a group of negative ions. In practice, the ligands may be polar molecules, the negative ends of which generate a negative field. Therefore, as it is applied to the study of bonding in coordination compounds, the term ligand field is appropriate.

Coordination complexes are composed of a metal ion held to a number of ligands by coordinate bonds. At this point the ligands will be considered to be interacting with the metal by predominantly electrostatic forces, and the ligand field model is adaptable to bonding arrangements in such complexes. Before the effects caused by the ligands can be understood, it is necessary to have a clear picture of the spatial arrangement of the  $5d$  orbitals, regardless of whether these are  $3d$ ,  $4d$ , or  $5d$  orbitals, that are oriented as shown in Figure 20.2. Keep in mind that the signs on the lobes of the orbitals are *mathematical* signs.

### 20.7.1 Octahedral Fields

In a “free” or isolated gaseous ion, the five  $d$  orbitals are degenerate. If the metal is surrounded with an electrostatic field that is identical in all directions, the  $d$  orbitals will all be affected to the same extent. Such a field is called a spherically symmetric field. Because this field is generated by a *negative* charge, there is a repulsion between the metal orbitals and the negative field. As a result of this repulsion, all the orbitals are raised in energy, but because the field is the same in all directions, there is no preferred direction so all are raised by an equal amount as shown in Figure 20.8.

Consider a complex containing a metal ion surrounded by six ligands in a regular octahedral arrangement as shown in Figure 20.9 where the ligands lie along the  $x$ ,  $y$ , and  $z$  axes. Because the lobes of the  $d_{x^2-y^2}$  and  $d_{z^2}$  orbitals lie *along* these axes, they are directed toward the six ligands. They will, therefore, experience greater repulsion than the orbitals that have lobes directed *between* the axes and, hence, between the ligands. Accordingly, the  $d_{xy}$ ,  $d_{yz}$ , and  $d_{xz}$  orbitals experience less repulsion than the  $d_{x^2-y^2}$  and  $d_{z^2}$  orbitals. As a result of the difference in repulsion, the five  $d$  orbitals no



**FIGURE 20.8** The energies of a set of five  $d$  orbitals in a free gaseous ion and with the ion surrounded by a field of negative charge.

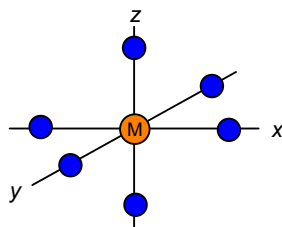


FIGURE 20.9 An octahedral complex  $ML_6$  with six ligands located on the  $x$ ,  $y$ , and  $z$  axes.

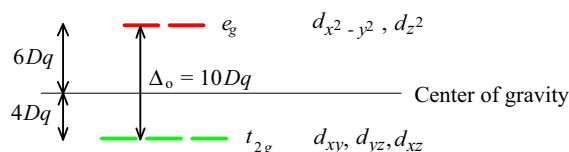
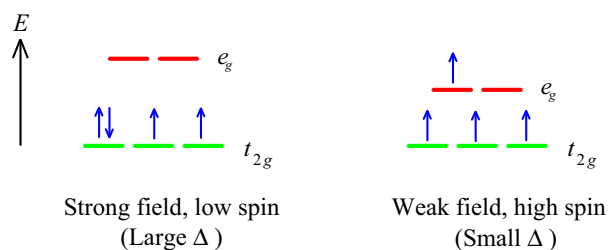


FIGURE 20.10 Splitting of a set of  $d$  orbital in an octahedral field.

longer have the same energy so this difference in repulsion causes the  $d$  orbitals to be split into two groups. The two orbitals having higher energy comprise one set and the three orbitals having lower energy than the other set. Because the groups of orbitals do not have the same energy they are designated differently with the orbitals of higher energy called the  $e_g$  set and those of lower energy the  $t_{2g}$  set. Note that the energies of *all* the orbitals are raised relative to those in the isolated ion because of repulsion, but some are raised more than others. This splitting of the  $d$  orbitals in the octahedral field is called the *ligand field splitting*. It is ordinarily referred to as  $\Delta_o$  or  $10 Dq$ , where  $Dq$  is simply an energy unit that is determined by the nature of complex. The amount of energy by which two orbitals are raised in energy must be equaled by the amount the other three are lowered in energy because there is no other energy change. Therefore, with respect to energy, a “center of gravity” or *barycenter* is maintained. This situation is shown in Figure 20.10.

It is now easy to see why an ion having a  $d^4$  configuration can be found either with all of the electrons unpaired or with only two of them unpaired. If the ligand field splitting is large enough to overcome electron–electron repulsion, two of the electrons will be forced to occupy one of the  $t_{2g}$  orbitals so there must be one electron pair. If the ligand field splitting is small, electron pairing cannot be forced, and the four electrons will be unpaired with three in the  $t_{2g}$  set and one in the  $e_g$  set. These situations can be shown as follows for a  $d^4$  ion.



However, it is still necessary to understand what factors determine the magnitude of the ligand field splitting energy so that predictions can be made as to the nature of complexes formed between specific metal ions and ligands. As described above, the magnetic moments of the complexes can be used to determine how the electrons are arranged because one case has four unpaired electrons but the other has only two.

## 20.7.2 Tetrahedral, Tetragonal, and Square Planar Fields

The splitting of the  $d$  orbitals described above is dependent on where the ligands are located around the metal. In other words, the ordering of the  $d$  orbitals with regard to energy will be different for complexes having different structures. Therefore, the arrangements of the  $d$  orbitals in complexes having structures other than octahedral will now be described. Let us first consider a tetrahedral complex that has the structure shown in Figure 20.11.

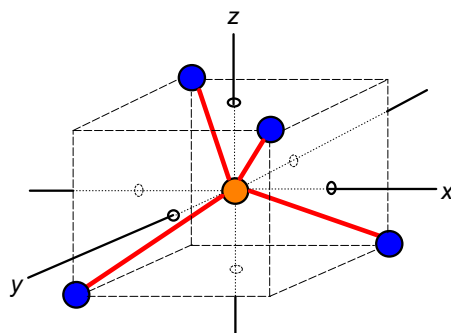
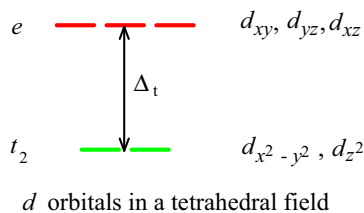


FIGURE 20.11 The structure of a tetrahedral complex.

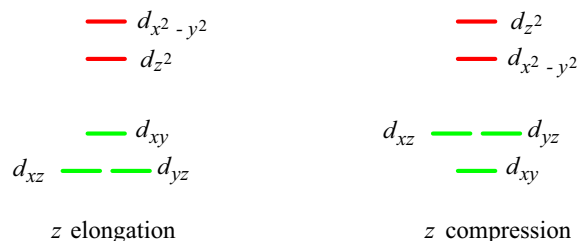
In a tetrahedral complex, the ligands are located *between* the axes and none of the  $d$  orbitals point directly at the ligands, but the  $d_{xy}$ ,  $d_{yz}$ , and  $d_{xz}$  orbitals are pointed more closely toward the ligands than are the  $d_{x^2-y^2}$  and  $d_{z^2}$  orbitals. Consequently, the  $d_{xy}$ ,  $d_{yz}$ , and  $d_{xz}$  orbitals experience greater repulsion and lie higher in energy than the  $d_{x^2-y^2}$  and  $d_{z^2}$  orbitals as shown below.



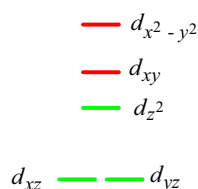
It should be noted that the triply degenerate set of orbitals is labeled as  $t_2$  with no  $g$  subscript. The  $g$  subscript refers to *gerade* or “even” and refers to preservation of sign on reflection through a center of symmetry. Because there is no center of symmetry for a tetrahedron (see Chapter 3), the  $g$  subscript is dropped giving the  $t_2$  and  $e$  designations. From the figure, it should be clear that because none of the orbitals point directly at the ligands, the splitting of the orbitals in a tetrahedral complex,  $\Delta_t$ , will not be as great as it is in the octahedral case. Also, in a tetrahedral complex there are only four ligands causing the splitting instead of six as in an octahedral complex. These factors cause the energy separating the two groups of orbitals in a tetrahedral complex to be much smaller than that occurring in an octahedral complex formed from the same metal and same ligands. In fact, it can be shown that if the same metal and same ligands were to form tetrahedral and octahedral complexes having identical metal–ligand bond distances, the relationship would be  $\Delta_t = (4/9)\Delta_o$ .

A complex containing six ligands but with those on the  $z$ -axis lying farther away from the metal ion than the other four is said to have a *tetragonal distortion*. Actually, the distance along the  $z$ -axis may be either greater or smaller than the others and the distortion is still called a tetragonal distortion. One case is referred to as  $z$  *elongation* and the other is called  $z$  *compression*. The effects of these distortions on the  $d$  orbitals are easy to see.

Consider an octahedral complex in which the ligands on the  $z$ -axis are at a greater distance from the metal than are those lying along the  $x$  and  $y$  axes. In this case, there will be less repulsion of the orbitals that project in the  $z$ -direction. The greatest effect will be on the  $d_{z^2}$  orbital that points directly at the ligands on the  $z$ -axis. However, in order to preserve the barycenter of the  $t_{2g}$  and  $e_g$  sets of orbitals, it is necessary for the  $d_{x^2-y^2}$  orbital to be raised in energy to offset the decrease in energy of the  $d_{z^2}$  orbital. Moreover, the  $d_{xz}$  and  $d_{yz}$  orbitals will also experience less repulsion than if the complex had regular octahedral geometry because they have a  $z$ -directional component. Consequently, the  $d_{xy}$  orbital must increase in energy to offset the decrease in energy of the  $d_{xz}$  and  $d_{yz}$  orbitals. As a result of these changes in orbital energies, the ordering of the  $d$  orbitals is altered to give the arrangement shown in the figure on page 338. If the ligands on the  $z$ -axis are moved closer to the metal ion than are those on the  $x$  and  $y$  axes, the orbitals projecting in the  $z$  direction are increased in energy and those that do not are correspondingly lowered in energy. There will be an inversion of each group of orbitals with regard to the  $z$  elongation case. Shown below are the arrangements of  $d$  orbitals for these tetragonally distorted cases.



A square planar field generated by four ligands lying on the  $x$ - and  $y$ -axes can be considered as an extreme case  $z$  elongation in which the ligands have been removed to an infinite distance from the metal ion. The effects on the  $d$  orbitals are the same as those described for  $z$  elongation but they occur to a considerably greater extent. In that case, the  $d_{z^2}$  orbital has its energy lowered to such an extent that it lies below the  $d_{xy}$  orbital. Consequently, in a square planar field, the  $d$  orbitals are arranged as shown below. It can be shown that the difference in energy between the  $d_{xy}$  and the  $d_{x^2-y^2}$  orbitals is exactly the same as  $\Delta_o$  if the complex contains the same metal and ligands in both cases.

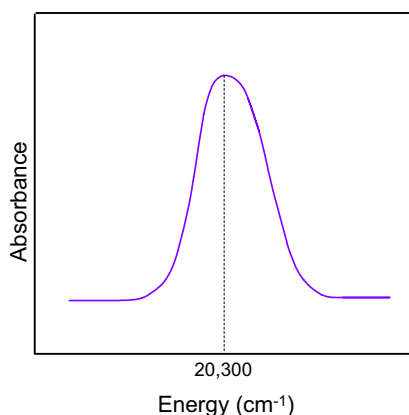


Splitting of  $d$  orbitals in a square planar field

Throughout this discussion, the splitting of the  $d$  orbitals caused by ligand fields has been shown. When there are electrons in the  $d$  orbitals, it is possible to induce transitions spectroscopically. The electron transitions that occur in a given complex are determined by the differences between energies of the  $d$  orbitals and, therefore, spectral studies provide a way to determine these energies experimentally. Although approximate values for  $\Delta_o$  can be calculated, this parameter is best considered as an experimental value to be determined, usually by spectroscopic methods. A discussion of the interpretation of spectra to determine ligand field splittings is beyond the scope of this book.

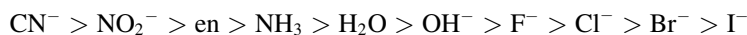
### 20.7.3 Factors Affecting $\Delta$

The extent to which the set of  $d$  orbitals is split in the electrostatic field produced by the ligands depends upon several factors. Two of the most important factors are the nature of the ligands and the nature of the metal ion. In order to see this effect, consider the complex ion  $\text{Ti}(\text{H}_2\text{O})_6^{3+}$ . The  $\text{Ti}^{3+}$  ion has a single electron in the  $3d$  level, and we refer to it as a  $d^1$  ion. In the octahedral field generated by six  $\text{H}_2\text{O}$  molecules, the single electron will reside in one of the three degenerate  $t_{2g}$  orbitals. Under spectral excitation, the electron is promoted to an  $e_g$  orbital giving rise to an absorption spectrum consisting of a single peak that can be represented as shown in Figure 20.12.



**FIGURE 20.12** Absorption spectrum for the complex  $[\text{Ti}(\text{H}_2\text{O})_6]^{3+}$ .

The maximum in the spectrum for  $[\text{Ti}(\text{H}_2\text{O})_6]^{3+}$  occurs at  $20,300\text{ cm}^{-1}$  which is equivalent to  $243\text{ kJ mol}^{-1}$ . This gives the value of  $\Delta_o$  directly, but only the case of a  $d^1$  ion is this simple. Other complexes containing the  $\text{Ti}^{3+}$  ion (e.g.,  $[\text{Ti}(\text{NH}_3)_6]^{3+}$ ,  $[\text{TiF}_6]^{3-}$ ,  $[\text{Ti}(\text{CN})_6]^{3-}$ , etc.) could also be prepared and spectra obtained for these complexes. If this were done, it would be observed that the absorption maximum occurs at a different energy for each complex. Because that maximum corresponds to the splitting of the  $d$  orbitals, the ligands could be ranked in terms of their ability to cause the splitting of orbital energies. Such a ranking is known as the *spectrochemical series*, and for several common ligands the following order of decreasing energy is observed.



Although the ideas of ligand field theory have been illustrated in this situation using the simplest ( $d^1$ ) case, it is possible to extend the spectroscopic study of metal complexes to include those that have other numbers of  $d$  electrons. Based on the study of many complexes, some generalizations about the ligand field splitting can be summarized as follows:

1. In general, the splitting in tetrahedral fields is only about half as large as that in octahedral fields.
2. For divalent ions of first row transition metals, the aqua complexes give splittings of about  $7000\text{--}14,000\text{ cm}^{-1}$ .
3. For complexes of +3 metal ions, the value of  $\Delta_o$  is usually about 30–40% larger than for the same metal ion in the +2 state.
4. For metal ions in the second transition series, the splitting is usually about 30% larger than for the first row metal ion having the same number of electrons in the  $d$  orbitals and the same charge. A similar difference exists between the third and second transition series.

Because of the effects described above, many ligands that give high spin complexes with first row transition metals give low spin complexes with second and third row metals. For example,  $[\text{NiCl}_4]^{2-}$  is high spin (tetrahedral), whereas  $[\text{PtCl}_4]^{2-}$  is low spin (square planar).

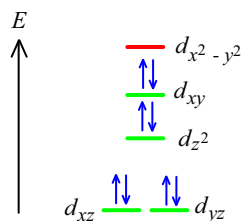
## 20.7.4 Ligand Field Stabilization Energy

If the splitting by the ligand field is large enough to cause pairing of electrons in a  $d^4$  ion, it is called a *strong field*. If pairing does not result, the field is called a *weak field*. It is easy to see that there is also another difference between the two cases. The  $e_g$  orbitals are *lowered* by  $4Dq$  but the  $t_{2g}$  orbitals are *raised* by  $6Dq$ . In the high spin case, there will be  $3 \times 4Dq = 12Dq$  units of energy gained by placing the three electrons in the  $t_{2g}$  orbitals but there will be  $6Dq$  sacrificed by placing the fourth electron in the  $e_g$  states. Thus, there is an overall energy gain of only  $6Dq$  units. In the low spin case, all four of the electrons reside in the  $t_{2g}$  orbitals resulting in a gain of  $4 \times 4Dq = 16Dq$  units of energy. The energy of stabilization resulting from populating the lower energy orbitals first is called the *ligand field stabilization energy* (LFSE). The procedure described here for a  $d^4$  ion could be carried out for other numbers of  $d$  electrons in both strong and weak fields. Table 20.5 shows the LFSE for the various numbers of  $d$  electrons in strong and weak field complexes.

**TABLE 20.5** Ligand Field Stabilization Energies in  $Dq$  Units

$d^n$	Weak Field	Strong Field
$d^1$	−4	−4
$d^2$	−8	−8
$d^3$	−12	−12
$d^4$	−6	−16
$d^5$	0	−20
$d^6$	−4	−24
$d^7$	−8	−18
$d^8$	−12	−12
$d^9$	−6	−6
$d^{10}$	0	0

The effects of the magnitude of  $\Delta$  can be seen in several ways. The most obvious is whether a particular complex is high spin or low spin. If a complex containing  $\text{Ni}^{2+}$  that has a  $d^8$  configuration is considered, it is found that complexes containing four ligands may be either tetrahedral or square planar depending on what the ligands are. The tetrahedral arrangement can be explained by assuming that none of the  $d$  orbitals is vacant (electron pairing does not occur) so that the hybridization of the metal ion orbitals is  $sp^3$ . For a square planar arrangement, the hybrid orbital type is  $dsp^2$ . One  $d$  orbital must be vacated by forcing the pairing of the eight electrons in the other four orbitals. In the square planar field, the  $d$  orbitals are arranged as shown below.

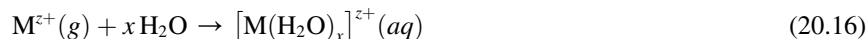


The  $d$  orbitals in a square planar field

Whether or not electron pairing will occur depends on the magnitude of the splitting of the  $d$  states. If the energy difference is greater than the pairing energy, the electrons will populate the levels as shown above and  $dsp^2$  hybridization is possible. For some ligands, the splitting is not large enough to force the pairing so a tetrahedral arrangement of the four ligands results. Consequently, complexes as  $[\text{Ni}(\text{NH}_3)_4]^{2+}$ ,  $[\text{Ni}(\text{H}_2\text{O})_4]^{2+}$ , and  $[\text{NiCl}_4]^{2-}$  are tetrahedral. Complexes of  $\text{Ni}^{2+}$  with ligands such as  $\text{CN}^-$  or  $\text{NO}_2^-$  are square planar as expected based on the positions of these ligands in the spectrochemical series.

The second and third row  $d^8$  ions  $\text{Pd}^{2+}$  and  $\text{Pt}^{2+}$  have considerably larger splitting of the  $d$  states than does  $\text{Ni}^{2+}$ . Consequently, the splitting is so large for these ions that only square planar complexes result, regardless of where the ligands fall in the spectrochemical series, and there are no known tetrahedral complexes of  $\text{Pd}^{2+}$  and  $\text{Pt}^{2+}$ .

Another manifestation of LFSE can be seen from the heats of hydration of the transition metal ions. For example, the hydration of a gaseous ion results in the formation of an aqua complex as represented by the equation



If  $\text{M}^{z+}$  is not a transition metal ion, the enthalpy of hydration will be determined by the size and charge of the ion. However, when a transition metal ion,  $\text{M}^{z+}$ , forms an aqua complex, not only are the size and charge of the ion important, but also there is an energy associated with electrons populating the  $d$  orbitals which are split by the ligand field. Thus, in addition to the usual electrostatic energy associated with the hydration of a cation, there will also be the LFSE released as the ion is solvated. The following equation represents the hydration process.



The enthalpy of hydration of  $+2$  metal ions is usually in the range of  $1600\text{--}2500 \text{ kJ mol}^{-1}$ . Because the sizes of the metal ions decrease in going to the right in the transition series, there will naturally be a general increase in hydration enthalpy in progressing from  $\text{Ca}^{2+}$  to  $\text{Zn}^{2+}$ . Figure 20.13 shows the hydration enthalpies for the  $+2$  ions of the first transition series when plotted in terms of the number of electrons in the  $d$  orbitals. The aqua complexes of these  $+2$  transition metal ions are all high spin (especially so for the  $+2$  ions where  $\Delta_o$  is small) so it is easy to see that for a  $d^1$  case there is an “extra” heat of  $4Dq$  released. For a  $d^2$  ion, the additional heat is  $8Dq$ , etc., as can be seen from the values shown in Table 20.5. For a high spin complex of a  $d^5$  ion (such as  $\text{Mn}^{2+}$ ), the LFSE is 0, and the enthalpy of hydration agrees well with that expected for an ion of that size and charge because there is no ligand field stabilization in that case.

It should be emphasized that although in principle one could determine  $Dq$  for aqua complexes in this way, it is highly impractical. First, hydration enthalpies of this magnitude are not known to high accuracy. Second, there is not a ready source of  $+2$  gaseous metal ions, and it would also involve determining small differences between large numbers, the heats of hydration of metal ions. As a result, this is not a practical type of measurement.

There are several other types of thermodynamic data that reflect the ligand field stabilization caused by splitting the  $d$  orbitals. For example, the lattice energies of the  $\text{MCl}_2$  (where  $\text{M}$  is a  $+2$  transition metal ion) compounds also show a “double-humped” shape when plotted as shown in Figure 20.8. However, these types of data will not be discussed because the trends follow naturally from the principles that have already been presented.



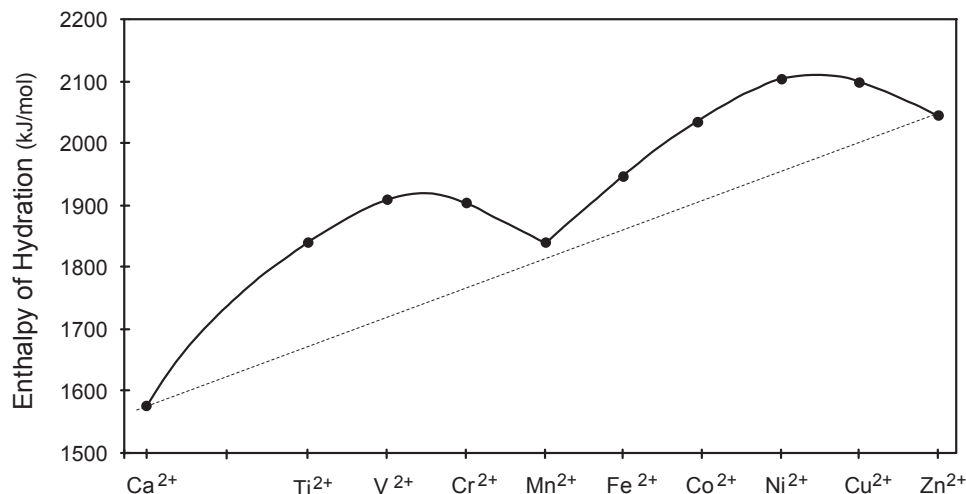


FIGURE 20.13 Enthalpies of hydration for some +2 metal ions of the first transition series.

## 20.8 JAHN–TELLER DISTORTION

One of the less obvious applications of ligand field theory is that it enables one to explain a curious fact. In solid  $\text{CuCl}_2$ , the  $\text{Cu}^{2+}$  ion is surrounded by six  $\text{Cl}^-$  ions. In the complex, it found that four  $\text{Cl}^-$  ions are at 230 pm from the copper ion and the other two are at a distance of 295 pm. The reason for this distortion from a regular octahedral structure lies in the way in which the  $d$  orbitals are populated. The  $\text{Cu}^{2+}$  ion has a  $d^9$  configuration with the orbitals having energies as shown in Figure 20.14 for a regular octahedral complex and a complex distorted along the  $z$ -axis.

When one examines the orbital splitting pattern for a complex having tetragonal distortion (elongation), it is seen that the subsets of  $d$  orbitals are split in energy. The splitting patterns showing the nine electrons populating the orbitals is illustrated in Figure 20.14. In the elongated complex, the energy of the  $d_{z^2}$  orbital is decreased by  $\delta_1$  but that of the  $d_{x^2-y^2}$  is increased by  $\delta_1$ . However, the  $d_{z^2}$  orbital is doubly occupied, whereas the  $d_{x^2-y^2}$  orbital contains only one electron. Therefore, there is a net decrease in the total energy of the orbitals by the amount  $\delta_1$  in the elongated structure. In other words, the complex will be more stable if the upper set ( $e_g$ ) of orbitals is split as shown in Figure 20.14(b). The reason for this is that the orbitals are originally degenerate, but they are unequally populated. Consequently, changing the energies of the two orbitals and placing the larger number of electrons in the one of lower energy will lead to a stabilization of the complex. A distortion of this type is known as a *Jahn–Teller distortion* (described in 1937 by H. A. Jahn and Edward Teller), and it occurs when orbitals of equal energy are unequally populated so that distorting the complex removes the orbital degeneracy.

Because of Jahn–Teller distortion, it is normally expected that octahedral complexes of  $\text{Cu}^{2+}$  will exhibit distortion. However, this is not the only case of this type. Note that a  $d^4$  high spin complex would have one electron in each of the four orbitals of lowest energy. That means that the orbitals obtained by splitting the upper states would be unequally populated because there is only one electron to be placed in the  $d_{z^2}$  orbital. Therefore, splitting the orbitals by distorting the octahedral structure would also lead to an overall reduction in energy in this case.

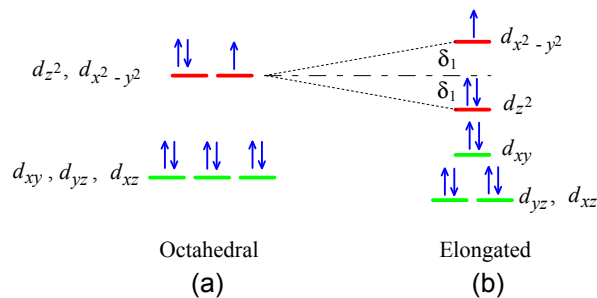


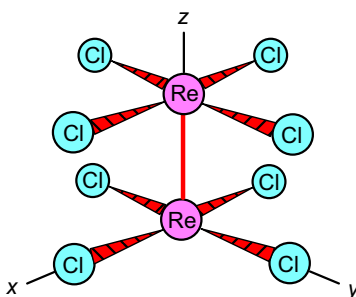
FIGURE 20.14 Population of  $d$  orbitals for a  $d^9$  complex with octahedral or distorted octahedral structure.

It is important to remember that the splitting of the  $t_{2g}$  orbitals also occurs, but the splitting is smaller than that of the  $e_g$  orbitals. Although it might be expected that complexes of  $d^1$  and  $d^2$  ions would undergo Jahn–Teller distortion, such distortion would be extremely small. In fact, there are some other problems in studying this type of distortion because of the short lifetimes of excited states and rearrangement of the complexes.

## 20.9 COMPLEXES HAVING METAL–METAL BONDS

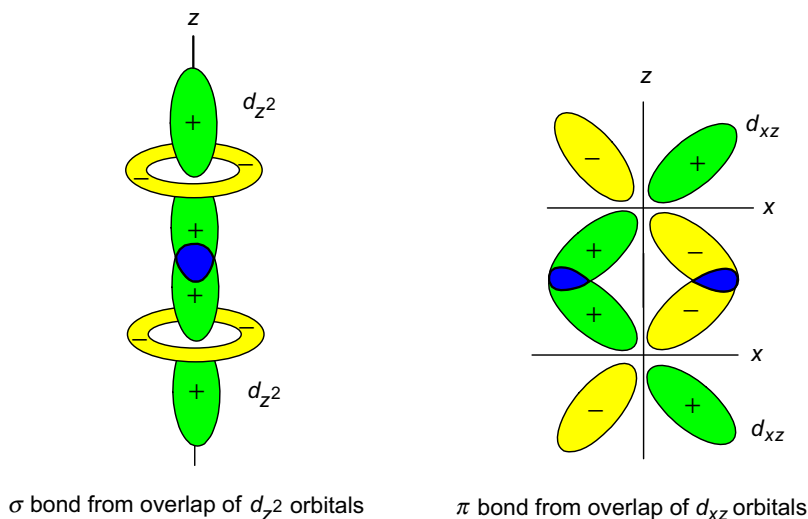
In addition to metal complexes that have coordinate bonds formed between metals and electron pair donors there are many others that also have bonds between metal atoms. One such complex is  $[\text{Re}_2\text{Cl}_8]^{2-}$  that has the bonding arrangement that can be shown as  $[\text{Cl}_4\text{Re}-\text{ReCl}_4]^{2-}$ . In this ion, the Re–Re bond length measures only 224 pm in length, which is indicative of a very short and strong bond. The elucidation and interpretation of the structure provide some interesting insight into the nature of bonds between metal atoms.

The very short bond between the metal atoms has been interpreted as arising from the interaction of multiple orbitals on the two atoms. First, the metal atoms are placed on the  $z$ -axis and the chlorine atoms lie on the  $x$ - and  $y$ -axes in each “half” of the structure, which can be shown as follows.

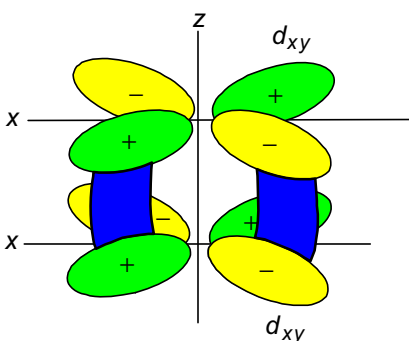


Therefore, the  $d_z^2$  orbitals on the metal atoms can interact in one region of space with the formation of a  $\sigma$  bond. However, the  $d_{xz}$  and  $d_{yz}$  orbitals on the metal atoms can overlap in two regions of space to form two  $\pi$  bonds. These orbital interactions can be shown as illustrated in Figure 20.15.

There is yet another possibility for forming an additional bond *between* the two metal atoms. The chlorine atoms lie on the  $x$ - and  $y$ -axes on either end of the structure. Therefore, the  $d_{xy}$  orbitals lie *between* the chlorine atoms so that a weak combination of these orbitals on each of the Re atoms can result in overlap of the four lobes of that orbital on each “end” of the structure. Overlap of the two  $d_{xy}$  orbitals in four regions produces a  $\delta$  bond. The orientation of the  $d_{xy}$  orbitals is shown in Figure 20.16.



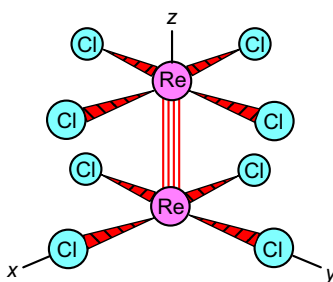
**FIGURE 20.15** Overlap of  $d_z^2$  orbitals to form a  $\sigma$  bond and overlap of  $d_{xz}$  orbitals to form a  $\pi$  bond in  $[\text{Re}_2\text{Cl}_8]^{2-}$ . Overlap of the  $d_{yz}$  orbitals results in a second  $\pi$  bond.



$\delta$  bond from overlap of  $d_{xy}$  orbitals

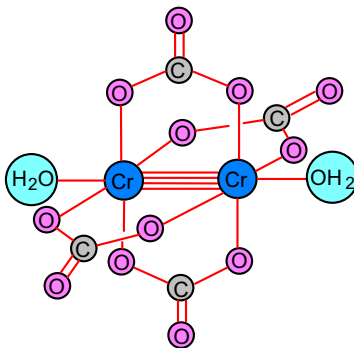
**FIGURE 20.16** Orientation of  $d_{xy}$  orbital that results in overlap in four regions of space giving rise to a  $\delta$  bond. Shading indicates overlap of only two of the four lobes.

Thus, the bonding in  $[\text{Re}_2\text{Cl}_8]^{2-}$  can be shown as



As a result of the very strong bond between the Re atoms, it is possible to carry out reactions in which chlorine atoms are replaced without disruption of the structure.

Many other molecules and ions involve metal–metal bonds and some have been known for a very long time. For example, if a compound of Cr(III) is reduced by the reaction with zinc and HCl, the solution has a blue color due to the presence of hydrated  $\text{Cr}^{2+}$ . Passing that solution into one containing acetic acid results in the formation of a brick red precipitate that historically was called chromous acetate. However, more recent studies have shown that the precipitate is actually  $[\text{Cr}_2(\text{CH}_3\text{COO})_4(\text{H}_2\text{O})_2]$  in which there is a quadruple bond between the chromium atoms. The four acetate ions form bridges between the chromium atoms. It has also been shown that the compound  $[\text{Cr}_2(\text{CO})_4(\text{H}_2\text{O})_2]$  has the structure



Numerous other types of compounds contain metal–metal bonds, notably some of the metal carbonyls (see Section 22.5).

In this chapter, an overview has been presented of several important aspects of the chemistry of coordination compounds. In addition to the elementary ideas related to bonding presented here, there is an extensive application of molecular orbital concepts to coordination chemistry. However, most aspects of the chemistry of coordination compounds treated in this book do not require this approach so it is left to more advanced texts. The references at the end of this chapter should be consulted for more details on bonding in complexes.

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## PROBLEMS

- Name the following compounds.
  - $[\text{Co}(\text{CN})_2(\text{NH}_3)_4]\text{NO}_3$
  - $\text{K}_2[\text{Cr}(\text{NCS})_2(\text{CN})_3\text{H}_2\text{O}]$
  - $[\text{Fe}(\text{NH}_3)_3(\text{H}_2\text{O})_3][\text{Pt}(\text{CN})_4]$
- Name the following species.
  - $(\text{NH}_4)_2[\text{Pd}(\text{SCN})_4]$
  - $[\text{Pt}(\text{SCN})_4]^{2-}$
  - $[\text{Mn}(\text{en})_2\text{Cl}_2]^+$
- Write formulas for the following compounds or ions.
  - Tetraamminebromochlorocobalt(III) nitrate
  - Hexaamminecobalt(III) hexacyanochromate(III)
  - Pentaamminepentacyano- $\mu$ -cyanodicobalt(III)
  - Hexa-*N*-thiocyanatochromate(III) ion
- Draw structures for each of the following species.
  - trans*-Bis(ethylenediamine)difluoroiron(III) ion
  - cis*-Diamminedi-*S*-thiocyanatoplatinum(II)
  - fac*-Triaquatricyanocobalt(III)
- Draw structures for each of the following species.
  - Diaquachlorotricyanoferrate(III) ion
  - Tris(oxalato)cobaltate(III) ion
  - Triaquacyanodipyridinechromium(III) ion
- Suppose a complex with the formula  $\text{ML}_3\text{X}_2$  where L and X are different ligands has a trigonal bipyramid structure. Sketch all the possible isomers for this complex.
- Write the formula or draw the structure as indicated.
  - The formula for a “polymerization” isomer of  $[\text{Pd}(\text{NH}_3)_2\text{Cl}_2]$
  - The structure of the *cis*-dichlorobis(oxalato)chromium(III) ion. Show all atoms.
  - The formula for a coordination isomer of  $[\text{Zn}(\text{NH}_3)_4][\text{Pd}(\text{NO}_2)_4]$
- Suppose a cubic complex has the formula  $\text{MX}_5\text{L}_3$  where L and X are different ligands. Sketch the structures for all the geometrical isomers possible.
- Explain why the solid compound  $\text{K}_3[\text{Ni}(\text{CN})_5]$  cannot be isolated.
- Describe the structure of  $\text{Cr}(\text{NH}_3)_6^{2+}$  in detail. Explain why the structure has the particular features that it does.
- For each of the following complexes, give the hybrid orbital type, the number of unpaired electrons, and the expected magnetic moment.
  - $\text{Co}(\text{H}_2\text{O})_6^{2+}$
  - $\text{FeCl}_6^{3-}$
  - $\text{PdCl}_4^{2-}$
  - $\text{Cr}(\text{H}_2\text{O})_6^{2+}$

12. For each of the following, sketch the  $d$  orbital energy level diagram and put electrons in the orbitals appropriately.
- $\text{Cr}(\text{CN})_6^{3-}$
  - $\text{FeF}_6^{3-}$
  - $\text{Co}(\text{NH}_3)_6^{3+}$
  - $\text{Ni}(\text{NO}_2)_4^{2-}$

For each of the cases above, give a specific ligand that would give a *different* arrangement of electrons with the same metal ion.

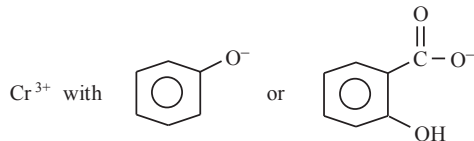
13. For each of the following, sketch the  $d$  orbital energy level diagram and put electrons in the orbitals appropriately.
- $\text{Mn}(\text{NO}_2)_6^{3-}$
  - $\text{Fe}(\text{H}_2\text{O})_6^{3+}$
  - $\text{Co}(\text{en})_3^{3+}$
  - $\text{NiCl}_4^{2-}$

For each of the cases above, give a specific ligand that would give a *different* arrangement of electrons with the same metal ion.

14. Explain why there are no low spin tetrahedral complexes.
15. Only one compound having the formula  $\text{Zn}(\text{py})_2\text{Cl}_2$  is known, but two different compounds are known having the formula  $\text{Pd}(\text{py})_2\text{Cl}_2$ . Explain these observations and give the hybrid orbital type and magnetic character of each complex.
16. Give a specific formula for each of the following.
- A ligand that will give a low spin complex with  $\text{Pd}^{2+}$  but not with  $\text{Ni}^{2+}$ .
  - The formulas for the first two known linkage isomers.
  - A ligand that is ambidentate.
  - A ligand that is hexadentate.
17. If it existed, what would be the approximate  $Dq$  value for the complex  $\text{Ti}(\text{H}_2\text{O})_4^{3+}$  (in  $\text{cm}^{-1}$ )? Briefly explain your answer.
18. The compound  $[\text{Co}(\text{NH}_3)_6]\text{Cl}_3$  is yellow-orange, whereas  $[\text{Co}(\text{H}_2\text{O})_3\text{F}_3]$  is blue. Explain why these materials are so different using the valence bond and ligand field theories.
19. Consider a linear complex with the two ligands on the  $z$ -axis. Sketch the energy level diagram for the  $d$  orbitals in this ligand field and label it completely.
20. Give the ligand field stabilization energy (in  $Dq$  units) for each of the following.
- $\text{CrF}_6^{3-}$
  - $\text{Fe}(\text{en})_3^{3+}$
  - $\text{CoBr}_4^{2-}$
  - $\text{Fe}(\text{NO}_2)_6^{4-}$
21. Give the ligand field stabilization energy (in  $Dq$  units) for each of the following.
- $\text{MnF}_6^{3-}$
  - $\text{Co}(\text{en})_3^{3+}$
  - $\text{CoCl}_4^{2-}$
  - $\text{VF}_6^{3-}$
22. Which of the following high spin complexes would you expect to undergo Jahn–Teller distortion?
- $\text{Cr}(\text{H}_2\text{O})_6^{2+}$
  - $\text{MnCl}_6^{3-}$
  - $\text{Fe}(\text{H}_2\text{O})_6^{3+}$
23. Which of the following high spin complexes would you expect to undergo Jahn–Teller distortion?
- $\text{Co}(\text{H}_2\text{O})_6^{2+}$
  - $\text{Cr}(\text{H}_2\text{O})_6^{3+}$
  - $\text{Ni}(\text{H}_2\text{O})_6^{2+}$

24. Predict which ligand of the pair given would give the more stable complex with the stated metal and justify your answer.

(a)



- (b) Hg<sup>2+</sup> with (C<sub>2</sub>H<sub>5</sub>)<sub>2</sub>O or (C<sub>2</sub>H<sub>5</sub>)<sub>2</sub>S  
 (c) Co<sup>3+</sup> with NH<sub>3</sub> or AsH<sub>3</sub>  
 (d) Zn<sup>2+</sup> with H<sub>2</sub>NCH<sub>2</sub>CH<sub>2</sub>NH<sub>2</sub> or H<sub>2</sub>NCH<sub>2</sub>CH<sub>2</sub>CH<sub>2</sub>NH<sub>2</sub>  
 (e) Ni<sup>0</sup> with PCl<sub>3</sub> or NH<sub>3</sub>
25. Although Co<sup>3+</sup> is too strong an oxidizing agent to exist in water, [Co(NH<sub>3</sub>)<sub>6</sub>]<sup>3+</sup> is stable in water. Explain this difference in the behavior of Co<sup>3+</sup>.
26. Explain why complexes that contain acetate ions as ligands are much less stable than those of the same metal ion when acetylacetonate ions are the ligands.
27. Predict the ways in which SO<sub>3</sub><sup>2-</sup> could coordinate to metal ions. If the other ligands are H<sub>2</sub>O, which way would it probably coordinate to Co<sup>3+</sup>? How would it likely coordinate to Pt<sup>2+</sup>?
28. Explain why high oxidation states of transition metals are stabilized by complexing the metal ions with NH<sub>3</sub>, whereas low oxidation states are stabilized by complexing with CO.
29. If a solid complex contained the linkage Cr<sup>3+</sup>–CN–Fe<sup>2+</sup> (only one of six bonds to each metal ion is shown), what would happen of the solid were heated? Explain your answer.
30. Suppose Ni<sup>4+</sup> formed a complex with NH<sub>3</sub>. What would you expect the magnetic moment of the complex to be? Explain your answer.

# Synthesis and Reactions of Coordination Compounds

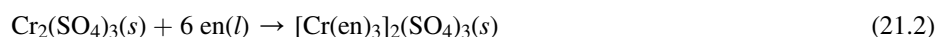
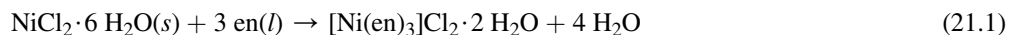
Coordination compounds are known to undergo an enormous number of reactions. However, the vast majority of these reactions can be classified into a rather small number of general types. The introduction to coordination chemistry presented in this book can describe only the most important types of reactions. The field of coordination chemistry is so vast that no single volume can provide complete coverage of the body of knowledge that exists concerning the reactions of the wide variety of complexes. In fact, the synthesis of coordination compounds and the reactions they undergo provide the basis for thousands of articles in journals each year. The treatment of the subject provided here gives a basis for further study, and the references cited at the end of this chapter provide greater depth and references to the original literature. It should be mentioned that the premier reference work on reactions of complexes in solution is the book by F. Basolo and R. G. Pearson, *Mechanisms of Inorganic Reactions* (see reference list).

## 21.1 SYNTHESIS OF COORDINATION COMPOUNDS

Because of the large number of types of coordination compounds, it should not be surprising to find that a wide variety of synthetic methods are used in their preparation. A brief survey of several of the widely used methods will be presented here, but others used in preparing complexes of olefins and metal carbonyls will be described in Chapter 22. The suggested readings at the end of this chapter give a more detailed treatment of this topic.

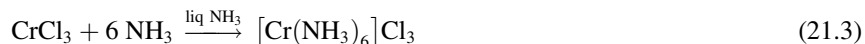
### 21.1.1 Reaction of a Metal Salt with a Ligand

The number of synthetic reactions of this type is enormous and the ligands include gases such as CO and liquids such as ethylenediamine (en).

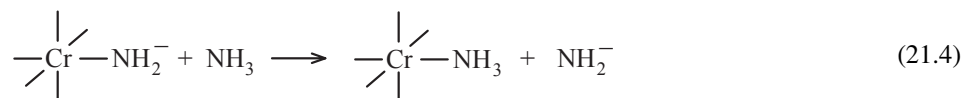


Because the solid  $[\text{Cr}(\text{en})_3]_2(\text{SO}_4)_3$  produced in the second reaction is formed as a hard cake, a novel technique can be used to produce a finely divided product that is easier to work with. If the ethylenediamine is dissolved in xylene and the solid  $\text{Cr}_2(\text{SO}_4)_3$  is added slowly to the refluxing mixture, the complex is obtained as small crystals that are insoluble in xylene and easily removed by filtration. This is an example of how a change in the reaction medium, not the reactants, leads to an improved synthesis.

The reactions between metal salts and ligands are not confined to aqueous solutions. For example, the preparation of  $[\text{Cr}(\text{NH}_3)_6]\text{Cl}_3$  can be carried out in liquid ammonia.



This process is very slow unless a small amount of sodium amide,  $\text{NaNH}_2$ , is added to catalyze the reaction. It is believed that the  $\text{NH}_2^-$  ion replaces  $\text{Cl}^-$  from  $\text{Cr}^{3+}$  more rapidly because it is a stronger nucleophile. After it has become coordinated, the  $\text{NH}_2^-$  can abstract  $\text{H}^+$  from a solvent molecule to regenerate  $\text{NH}_3$ .

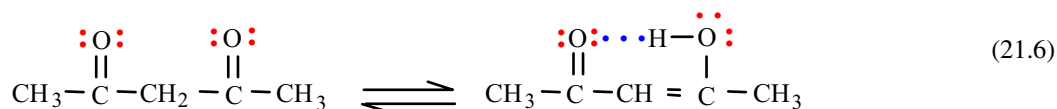


Hydroxocomplexes behave similarly in aqueous solutions because of the facile transfer of  $\text{H}^+$  between the solvent and the coordinated  $\text{OH}^-$  group.

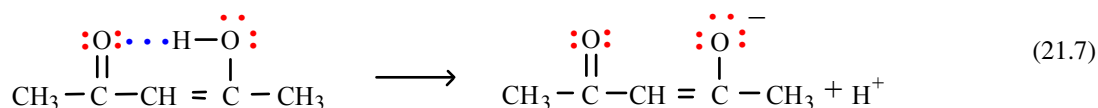
The reactions of butadiene,  $C_4H_6$ , with certain metal compounds are particularly interesting because bridged complexes are produced. This can occur because butadiene has two double bonds that can function simultaneously as electron pair donors to two metal atoms. When the reaction is carried out using  $CuCl$  and liquid butadiene at  $-10\text{ }^\circ\text{C}$ , the reaction can be represented by the following equation.



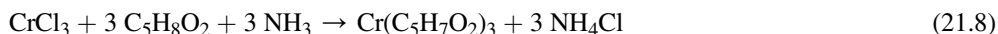
The synthesis of tris(acetylacetonate)chromium(III) shows additional features of this type of synthesis involving a metal compound reacting with a ligand. Acetylacetonate (2,4-pentadione),  $C_5H_8O_2$ , undergoes a tautomerization,



Further, the enol form loses the proton on the OH group easily in a reaction that is pH dependent. The loss of  $H^+$  can be shown as



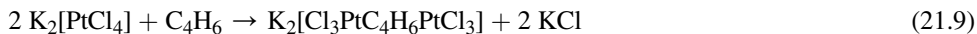
If  $NH_3$  is present, the acetylacetonate ion, acac, coordinates readily owing to the increased concentration of the anion and the stability of the complexes containing the anion. The reaction with  $CrCl_3$  is illustrative of this process.



### 21.1.2 Ligand Replacement Reactions

The replacement of one ligand in a complex by another is a very common type of reaction. These reactions (discussed in greater detail later in this chapter) form the basis for the synthesis of many types of complexes. The reactions may be carried out in aqueous solutions, nonaqueous solutions, or in the neat liquid and gas phases. Some illustrative examples will be described here.

A replacement reaction that produces a bridged complex containing butadiene occurs when gaseous butadiene reacts with a solution of  $K_2[\text{PtCl}_4]$ . Butadiene has two double bonds so it can function as a bridging ligand by bonding to two metals simultaneously.



Replacement reactions are important because frequently one type of complex is easily prepared and then it can be converted to another that cannot be obtained easily by a direct reaction. A reaction of this type is the following.



It should be noted that in this case a soft ligand, CO, is being replaced with another soft ligand that has phosphorus as the electron pair donor.

### 21.1.3 Reaction of Two Metal Compounds

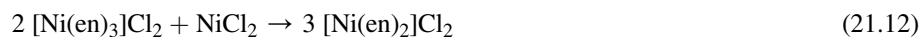
In some cases, it is possible to carry out a reaction of two metal compounds directly to prepare a complex. One such reaction is the following.



This reaction can be carried out in a variety of ways, but one of the most recently discovered is the use of ultrasound applied to a suspension of the solid reactants in dodecane. The ultrasonic vibrations cause cavities in the liquid which implode causing the particles of the solid to be driven together at high velocity causing them to react.

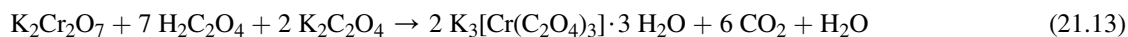


A variation of the reaction of two metal compounds occurs when one metal complex already containing ligands reacts with a simple metal salt to give a redistribution of the ligands. For example, such a reaction takes place between  $[\text{Ni}(\text{en})_3]\text{Cl}_2$  and  $\text{NiCl}_2$ .

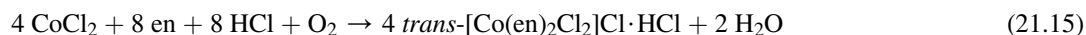
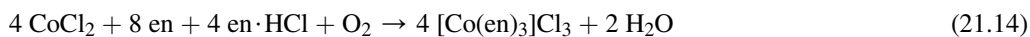


### 21.1.4 Oxidation–Reduction Reactions

A great number of coordination compounds are prepared by the oxidation or reduction of a metal compound in the presence of a ligand, and this technique has been used for a very long time. In some cases, the same compound serves as the source of the ligand and the reducing agent. Oxalic acid,  $\text{H}_2\text{C}_2\text{O}_4$ , is one such compound, and the preparation of the tris(oxalato) complex with  $\text{Co}^{3+}$  is typical of this behavior. The overall reaction is represented by the following equation.



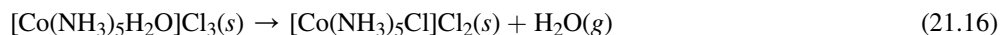
When oxidation of the metal is involved, a variety of oxidizing agents may be used in specific cases. Numerous complexes of  $\text{Co}(\text{III})$  have been prepared by the oxidation of solutions containing  $\text{Co}(\text{II})$ . In the following reactions, en is ethylenediamine.



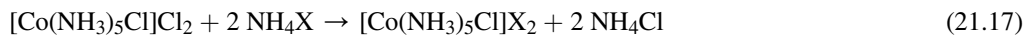
Heating the product in the second reaction yields *trans*- $[\text{Co}(\text{en})_2\text{Cl}_2]\text{Cl}$  by the loss of HCl. Dissolving that compound in water and evaporating the solution by heating results in the formation of *cis*- $[\text{Co}(\text{en})_2\text{Cl}_2]\text{Cl}$ .

### 21.1.5 Partial Decomposition

When coordination compounds are heated in the solid state, it is not uncommon for some ligands, particularly volatile ones such as  $\text{H}_2\text{O}$  and  $\text{NH}_3$ , to be driven off. When that happens, other groups can enter the coordination sphere of the metal or perhaps change bonding mode. For example, the reaction

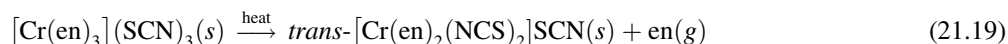
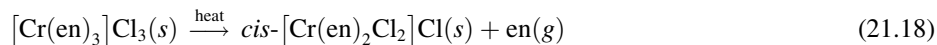


represents the usual way of preparing  $[\text{Co}(\text{NH}_3)_5\text{Cl}]\text{Cl}_2$ . The analogous compounds containing other anions (X) can be prepared by means of reactions such as



by mixing aqueous solutions of the reactants. The number of known reactions of this type is very large, and many of them have been studied in great detail from a kinetic standpoint.

Two additional preparative reactions that involve partial decomposition are the following.



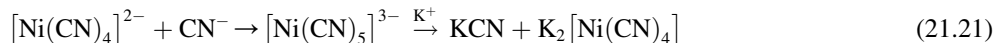
These reactions have been known for almost 100 years, and they have been extensively studied. The reactions are catalyzed by the corresponding ammonium salt in each case although other protonated amines can function as catalysts. It appears that the function of the catalyst is to react as a proton donor that facilitates the removal of an end of the coordinated ethylenediamine molecule from the metal.

### 21.1.6 Size and Solubility Relationships

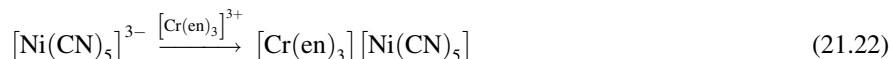
When an aqueous solution containing  $\text{Ni}^{2+}$  has a solution containing  $\text{CN}^-$  added to it,  $[\text{Ni}(\text{CN})_4]^{2-}$  is produced.



It is known that if  $\text{CN}^-$  is present in large excess, another cyanide ion adds to the  $[\text{Ni}(\text{CN})_4]^{2-}$  to form  $[\text{Ni}(\text{CN})_5]^{3-}$ . However, when  $\text{K}^+$  is added to the solution in an attempt to isolate solid  $\text{K}_3[\text{Ni}(\text{CN})_5]$ , the result is  $\text{K}_2[\text{Ni}(\text{CN})_4]$  and  $\text{KCN}$ .



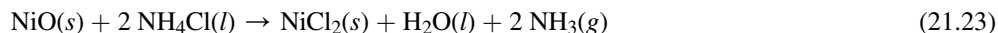
An enormously useful principle of chemistry states that species of similar electronic character (size and charge) interact best (see Chapter 6). For example,  $\text{Ba}^{2+}$  is precipitated as the sulfate in gravimetric analysis. Both are doubly charged ions of comparable size. This suggests that the best way to obtain a solid containing the  $[\text{Ni}(\text{CN})_5]^{3-}$  ion, a rather large ion of  $-3$  charge, would be to use a large  $+3$  cation. Such an ion is  $[\text{Cr}(\text{en})_3]^{3+}$ , and when a solution containing that ion is added to the solution containing  $[\text{Ni}(\text{CN})_5]^{3-}$ , solid  $[\text{Cr}(\text{en})_3][\text{Ni}(\text{CN})_5]$  is obtained.



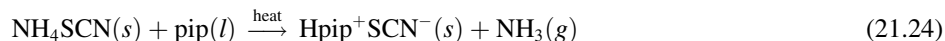
Many complex ions that form in solutions,  $[\text{Ni}(\text{CN})_5]^{3-}$  being one of them, are not particularly stable. Consequently, any technique that produces a more stable environment for such an ion becomes a factor to consider in synthesis. This case illustrates the importance of understanding general rules regarding the favorable interactions of species of similar size and charge.

### 21.1.7 Reactions of Metal Salts with Amine Salts

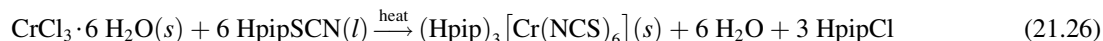
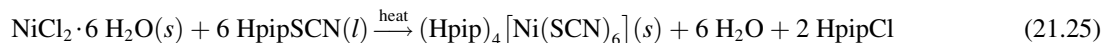
Amine salts such as  $\text{Hpy}^+\text{Cl}^-$  (where py represents pyridine) are acidic in the molten state as is  $\text{NH}_4\text{Cl}$ . Therefore, they will react with a variety of metal compounds, e.g., oxides, carbonates, etc., by acid–base reactions such as the following.



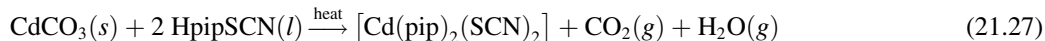
Hydrothiocyanic acid (also known as rhodanic acid in older terminology),  $\text{HSCN}$ , is a strong acid, so it is easy to prepare amine hydrothiocyanate salts such as that of piperidine,  $\text{C}_5\text{H}_{11}\text{N}$ , abbreviated pip. A simple synthesis of amine hydrothiocyanates utilizes the replacement of  $\text{NH}_3$  from  $\text{NH}_4\text{SCN}$  by the amine.



The product,  $\text{HpipSCN}$ , is known as piperidinium thiocyanate or piperidine hydrothiocyanate, and it has a melting point of  $95^\circ\text{C}$ . When this salt is heated to melting and compounds of transition metals are added, thiocyanate complexes of the metals result. For example, the following reactions can be carried out at  $100^\circ\text{C}$  when using a 10:1 ratio of the amine hydrothiocyanate to metal compound.



In some cases, such as the following, the metal complex contains both piperidine and thiocyanate as ligands.



Several reactions of this type have been carried out at  $0^\circ\text{C}$  by the application of ultrasound to mixtures of metal salts and  $\text{HpipSCN}$  suspended in dodecane. It appears that this method results in products of higher purity than when the molten salt is used. This is probably due to the fact that some of the products are not very stable at the temperature of the molten salt and mixtures result under those conditions, but there is no thermal decomposition when ultrasound is the source of energy at low temperature.

The synthetic reactions shown represent only a small fraction of the wide variety of techniques that have been developed for preparing complexes. A good starting point when searching for a method to prepare a specific type of complex is the 35 volume set known as *Inorganic Syntheses* (see references).

## 21.2 A SURVEY OF REACTION TYPES

Of all the types of reactions known for coordination compounds, the most thoroughly studied are those known as substitution reactions in which one ligand replaces another. Although other types of reactions will be described, much of the

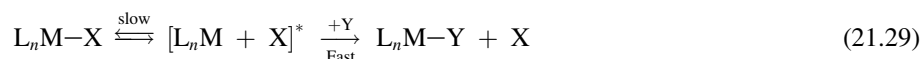
remainder of this chapter will be devoted to a description of substitution reactions from the standpoint of factors affecting their rates. However, before presenting a discussion of substitution reactions, a brief summary of several types of reactions will be given.

### 21.2.1 Ligand Substitution

As was mentioned earlier, the substitution of one ligand for another in a complex is the most common type of reaction that complexes undergo. If the starting complex contains  $n$  ligands of type L and one ligand X that is replaced by Y, this type of reaction can be shown in a general form by the equation



Although the subject of reaction mechanism will be considered later, it is appropriate at this time to mention that there are two possible limiting mechanisms for this type of reaction. In the first, the ligand X leaves slowly before Y enters so that the reaction can be represented by the following scheme.

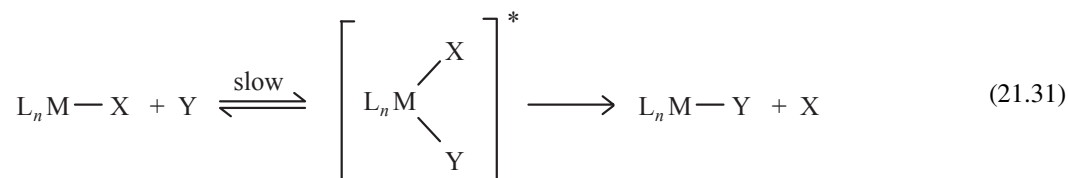


where L is a nonparticipating ligand and  $n$  is the number of ligands of that type. In this type of reaction, X is called the leaving group and Y is called the entering ligand. In this case, the transition state is denoted by  $[ ]^*$ , and it is the concentration of this species that determines the rate of reaction. The concentration of the transition state is determined by the concentration of the reacting complex  $L_nM-X$  so the rate of reaction is proportional to the concentration of that species. Thus, the rate law for the reaction can be written as

$$\text{Rate} = k_1[L_nM-X] \quad (21.30)$$

Such a mechanism is known as a *dissociative mechanism* because dissociation of the M-X bond is the slow (rate determining) step in the process. Such a mechanism can also be considered as a first-order or  $S_N1$  mechanism.

If the reaction takes place with Y entering *before* X leaves, it can be represented by the following scheme.



In this case, formation of the transition state requires both  $L_nM-X$  and Y, and the rate of the reaction depends on the concentration of both species. Therefore, the rate law can be written as

$$\text{Rate} = k_2[L_nM-X][Y] \quad (21.32)$$

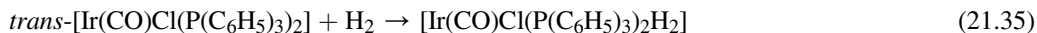
and the mechanism is called an *associative pathway*, but it is also described as second-order or  $S_N2$ . As will be shown when a more detailed discussion of substitution reactions is presented, many substitution reactions are not this simple, and the rate laws are frequently much more complicated than those shown above. Another aspect of substitution reactions that will be discussed is the enormous difference in the rates of the reactions. Some substitution reactions are so fast that they take place with rate constants as large as  $10^8 \text{ s}^{-1}$  or so slowly that the rate constants are as small as  $10^{-8} \text{ s}^{-1}$ , and some of the reasons for this difference will be explored later in this chapter.

### 21.2.2 Oxidative Addition (Oxad) Reactions

An oxidative addition reaction occurs when there is an increase in oxidation state of a metal that is accompanied by a simultaneous increase in its coordination number. Numerous examples of oxidative addition (sometimes referred to as *oxid*) reactions have already been presented for nonmetallic elements including the following.



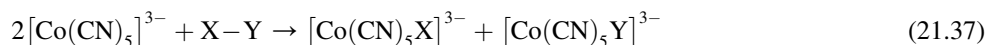
Each of these reactions involves an increase in the oxidation state of the central atom and the formation of additional bonds to that atom. In this chapter, the discussion of such oxad reactions will be limited to those involving transition metals in complexes. An interesting reaction of this type involving hydrogen is the following.



In this reaction, Ir is oxidized from the +1 oxidation state to +3 during the addition of a hydrogen molecule. When coordinated to the metal, hydrogen is considered to be coordinated as hydride ions,  $\text{H}^-$ . Thus, hydrogen has been added by increasing the coordination number of Ir to 6, and the metal has been oxidized to +3. A general oxad reaction can be represented as

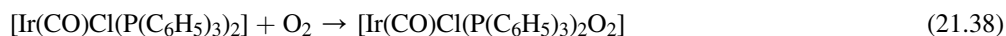


where  $\text{X}-\text{Y} = \text{H}_2, \text{Cl}_2, \text{HCl},$  or  $\text{CH}_3\text{I}$ . The reaction

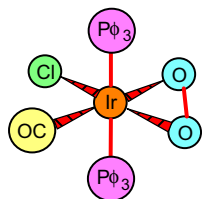


where  $\text{X}-\text{Y} = \text{H}_2, \text{Cl}_2, \text{H}_2\text{O}_2,$  or  $\text{CH}_3\text{I}$ , is of interest because  $[\text{Co}(\text{CN})_5]^{3-}$  is an important catalyst. In this case, two metal ions increase in oxidation state by one unit each instead of a single metal ion increasing by two units.

The addition of oxygen in the following oxad reaction is slightly different because when an oxygen molecule adds to the complex, the atoms do not become separated.



Therefore, the structure of the product (where  $\phi$  is the phenyl group,  $\text{C}_6\text{H}_5$ ) is



In this complex, the oxygen molecule behaves as if it is bonded to the metal in *cis* positions as a peroxide ion,  $\text{O}_2^{2-}$ .

By noting that the metal increases its oxidation state and coordination number simultaneously, it is possible to give the general requirements for oxidative addition to occur.

1. The metal must be able to change oxidation state by two units unless two metal ions are involved as in Eq. (21.37).
2. Two vacant sites are required where the entering ligands can bond to the metal. This often occurs as a square planar complex is converted to an octahedral one as shown in the following example.



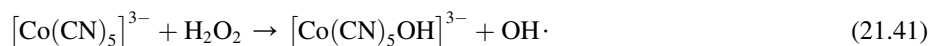
In this reaction,  $\text{Pt}^{2+}$  is oxidized to  $\text{Pt}^{4+}$  and the square planar  $[\text{PtCl}_4]^{2-}$  is converted into  $[\text{PtCl}_6]^{2-}$ , which has an octahedral structure.

3. The groups enter in *cis* positions if the atoms of the added molecule do not become separated before bonding to the metal.
4. In polar solvents, substances such as  $\text{HCl}$  are extensively ionized so it is possible for the H and Cl to enter separately at different times. Therefore, they may not enter in *cis* positions. However, if the reaction is carried out in a nonpolar solvent or in the gas phase, the attack is by an  $\text{HCl}$  molecule so addition takes place in *cis* positions.

Oxidative addition reactions do not all follow the same mechanism. For example, the reaction



is believed to follow a free radical mechanism in which the rate determining step can be represented as



Therefore, the reaction of  $[\text{Co}(\text{CN})_5]^{3-}$  with  $\text{H}_2$  illustrated in Eq. (21.37) follows the rate law

$$\text{Rate} = k[\text{H}_2][\text{Co}(\text{CN})_5^{3-}]^2 \quad (21.42)$$

Consequently, it appears to involve a concerted one-step mechanism. Breaking the H–H bond requires about  $435 \text{ kJ mol}^{-1}$  and this is compensated for by the formation of *two* Co–H bonds each of which releases about  $243 \text{ kJ mol}^{-1}$ .

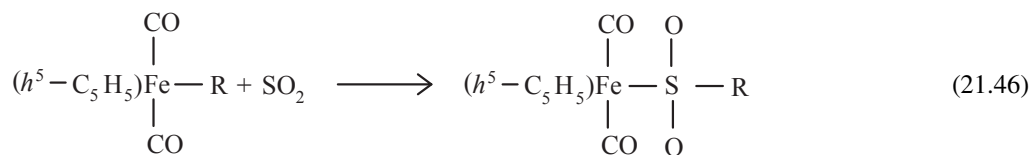
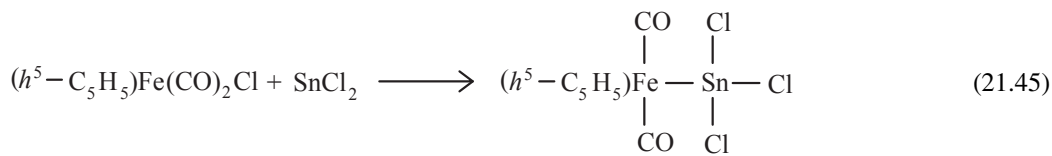
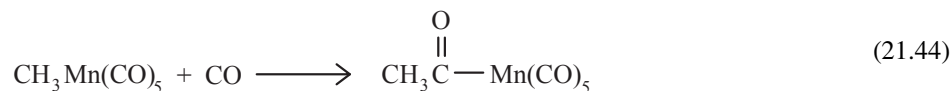
A *reductive elimination* reaction is one in which the oxidation state of the metal decreases by two units and the coordination number of the metal decreases. The effect of a reductive elimination reaction is the opposite of that produced by an oxidative addition.

### 21.2.3 Insertion Reactions

When a group enters a complex by being inserted between the metal and another ligand in the complex, the reaction is known as an *insertion reaction*. This type of reaction can be shown as follows.



Some molecules that will undergo insertion reactions are CO,  $\text{SnCl}_2$ ,  $\text{SO}_2$ ,  $\text{R}-\text{NC}$ ,  $\text{C}_2\text{H}_4$ , and  $(\text{CN})_2\text{C}=\text{C}(\text{CN})_2$ . The following equations show examples of insertion reactions.

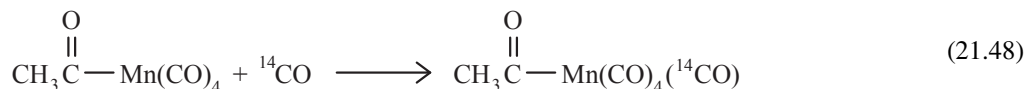


### 21.2.4 Group Transfer Reactions

A reaction in which a group is moved from a ligand to the metal or from the metal to a ligand is known as a *transfer reaction* (also referred to as a *migration reaction*). The reaction shown in Eq. (21.44) is an example of this type of reaction that has received a great deal of study. The insertion of CO in the complex  $\text{CH}_3\text{Mn}(\text{CO})_5$  does not involve simply sliding CO in the Mn–CH<sub>3</sub> bond. Mechanistic studies on this reaction have been carried out using  $^{14}\text{C}$  as the entering ligand followed by determining where the labeled  $^{14}\text{C}$  molecules reside in the complex. These studies show that the CO that becomes inserted between Mn and CH<sub>3</sub> is not the  $^{14}\text{C}$  that adds to the molecule. The first step involves one of the CO molecules coordinated elsewhere migrating from that position to interpose itself between the Mn and the CH<sub>3</sub>. This process is illustrated by the following equation.

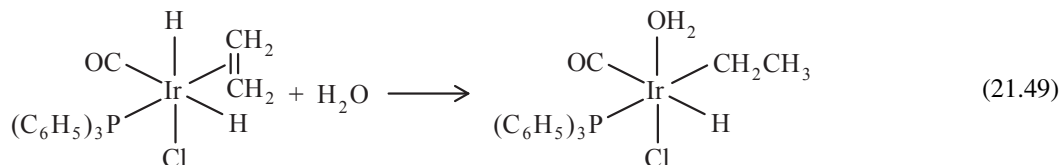


The labeled  $^{14}\text{C}$  now adds at the site vacated as CO migrates to the Mn–CH<sub>3</sub> bond to complete the coordination sphere.

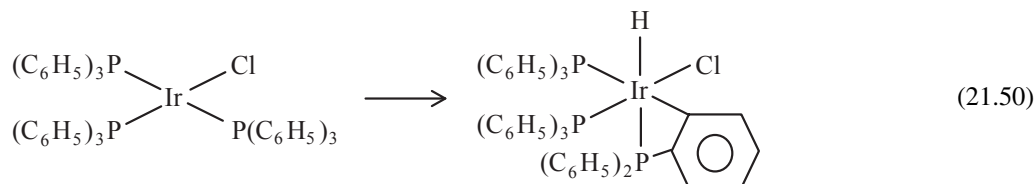


Therefore, this reaction is an *insertion reaction*, but it occurs by a mechanism that involves a *transfer* of a ligand from another position in the coordination sphere of the metal.

The transfer of hydrogen is believed to be one step in the hydrogenation of olefins in which complexes function as catalysts. The following equations represent examples of hydrogen transfer reactions.



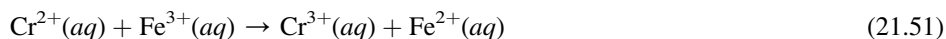
In this reaction, H is transferred to the olefin as H<sub>2</sub>O attaches at the vacated position. The reaction



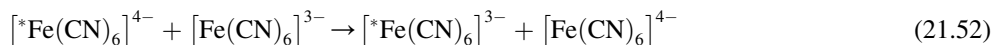
involves the transfer of hydrogen from one of the phenyl groups on P(C<sub>6</sub>H<sub>5</sub>)<sub>3</sub> to the metal as a bond forms between the metal and a carbon atom in the phenyl ring.

### 21.2.5 Electron Transfer Reactions

An aqueous solution containing complexes of two different metal ions may provide a medium for a redox reaction. In such cases, electrons are transferred from the metal ion being oxidized to the metal ion being reduced. The reaction between Cr<sup>2+</sup> and Fe<sup>3+</sup> in aqueous solution is an example of this type of reaction,



In this reaction, an electron is transferred from Cr<sup>2+</sup> to Fe<sup>3+</sup>, and such reactions are usually called *electron transfer* or *electron exchange* reactions. Electron transfer reactions may also occur in cases where only one type of metal ion is involved. For example, a reaction between two isotopes of Fe can be shown as

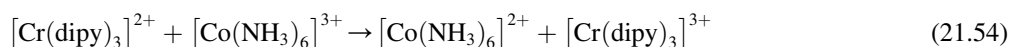
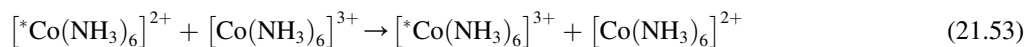


In this reaction, an electron is transferred from [\*Fe(CN)<sub>6</sub>]<sup>4-</sup> to Fe<sup>3+</sup> in [Fe(CN)<sub>6</sub>]<sup>3-</sup> (where Fe and \*Fe are different isotopes of iron). This is an electron transfer in which the product differs from the reactants only in that a different isotope of Fe is contained in the +2 and +3 oxidation states.

Electron transfer between metal ions contained in complexes can occur in two different ways, depending on the nature of the metal complexes present. If the complexes are inert, electron transfer occurring faster than the substitution processes must occur without breaking the bond between the metal and ligand. Such electron transfers are said to take place by an *outer sphere mechanism*. Thus, each metal ion remains attached to its original ligands and the electron is transferred through the coordination spheres of the metal ions.

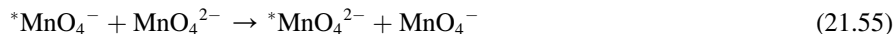
In the second case, the ligand replacement processes are more rapid than the electron transfer process. If this is the case (as it is with other complexes that undergo rapid substitution), a ligand may leave the coordination sphere of one of the metal ions and be replaced by forming a bridge utilizing a ligand already attached to a second metal ion. Electron transfer then occurs through a bridging ligand, and this is called an *inner sphere mechanism*.

For an outer sphere electron transfer, the coordination spheres of each complex ion remain intact. Thus, the transferred electron must pass through both coordination spheres. Reactions such as the following are of this type (where \* represents a different isotope).



In the reaction shown in Eq. (21.54), the d<sup>4</sup> complex containing Cr<sup>2+</sup> is inert owing to electron pairing caused by ligands giving the low spin state. There is an extreme variation in rates of electron transfer from very slow to very fast depending on the nature of the ligands present, and rate constants may vary from 10<sup>-6</sup> to 10<sup>8</sup> M<sup>-1</sup> s<sup>-1</sup>.

The electron exchange between manganate ( $\text{MnO}_4^{2-}$ ) and permanganate ( $\text{MnO}_4^-$ ) takes place in basic solutions as represented by the following equation.



In this reaction, \*Mn and Mn represent different isotopes of manganese. The rate law for this reaction can be written as

$$\text{Rate} = k[*\text{MnO}_4^-][\text{MnO}_4^{2-}] \quad (21.56)$$

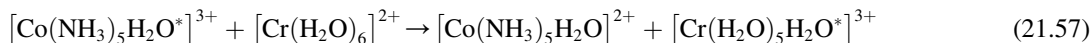
When the solvent contains  $\text{H}_2\text{O}^{18}$ , it is found that no  $^{18}\text{O}$  is incorporated in the  $\text{MnO}_4^-$  that is produced. Thus, the reaction is presumed to proceed without the formation of oxygen bridges. However, the nature of the cations present greatly affects the rate. The rate of the reaction varies with the cation present in the order  $\text{Cs}^+ > \text{K}^+ \approx \text{Na}^+ > \text{Li}^+$ . This supports the view that the transition state must involve a structure in which a cation is located between the  $\text{MnO}_4$  groups as shown below.



Presumably, the function of  $\text{M}^+$  is to “cushion” the repulsion of the two negative ions. The larger, softer  $\text{Cs}^+$  can do this more effectively than the smaller, harder ions such as  $\text{Li}^+$  or  $\text{Na}^+$ . Also, to form these bridged transition states, solvent molecules must be displaced from the solvation sphere of the cations. That process, would require more energy in the case of  $\text{Li}^+$  and  $\text{Na}^+$ , which are more strongly solvated as a result of their smaller size. When  $\text{Cs}^+$  ion is present (which forms effective bridges between the  $\text{MnO}_4$  units), the rate of electron exchange is linearly related to  $\text{Cs}^+$  concentration.

Similar results have been found for the electron exchange between  $[\text{Fe}(\text{CN})_6]^{3-}$  and  $[\text{Fe}(\text{CN})_6]^{4-}$ . In that case, the acceleratory effects are found to vary with the nature of the cation in the order  $\text{Cs}^+ > \text{Rb}^+ > \text{K}^+ \approx \text{NH}_4^+ > \text{Na}^+ > \text{Li}^+$  in accord with the size and solvation effects discussed above. For +2 ions, the order of effect on the rate is  $\text{Sr}^{2+} > \text{Ca}^{2+} > \text{Mg}^{2+}$ , which is also in accord with the softness of these species. Exchange in these outer sphere cases is believed to involve the formation of bridged species containing cations that are probably only partially solvated.

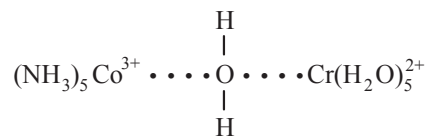
In aqueous solutions,  $\text{Cr}^{2+}$  is a strong reducing agent, and it reduces  $\text{Co}^{3+}$  to  $\text{Co}^{2+}$ , and a number of electron transfer reactions involving complexes of these metals have been studied. High spin complexes of  $\text{Cr}^{2+}$  ( $d^4$ ) are kinetically labile as are high spin complexes of  $\text{Co}^{2+}$  ( $d^7$ ). However, complexes of  $\text{Cr}^{3+}$  ( $d^3$ ) and low spin complexes of  $\text{Co}^{3+}$  ( $d^6$ ) are kinetically inert. For the exchange reaction shown as



(where  $\text{O}^*$  represents  $^{18}\text{O}$ ), it was found that the rate law is

$$\text{Rate} = k[\text{Co}(\text{NH}_3)_5\text{H}_2\text{O}^{*3+}][\text{Cr}(\text{H}_2\text{O})_6^{2+}] \quad (21.58)$$

It was also found that the  $\text{H}_2\text{O}^*$  is transferred quantitatively to the coordination sphere of  $\text{Cr}^{3+}$ . Thus, it appears that the electron is transferred from  $\text{Cr}^{2+}$  to  $\text{Co}^{3+}$  but the  $\text{H}_2\text{O}^*$  is transferred from  $\text{Co}^{3+}$  to  $\text{Cr}^{2+}$  as reduction occurs. As a result, electron transfer probably occurs through the bridged transition state that can be shown as



The  $\text{H}_2\text{O}$  forming the bridge then becomes part of the coordination sphere of the kinetically inert  $\text{Cr}^{3+}$ .

A large number of reactions similar to that shown in Eq. (21.57) have been studied in detail. One such reaction is



where X is an anion such as  $\text{F}^-$ ,  $\text{Cl}^-$ ,  $\text{Br}^-$ ,  $\text{I}^-$ ,  $\text{SCN}^-$ , or  $\text{N}^{3-}$ . The  $\text{Co}^{2+}$  produced is written as  $[\text{Co}(\text{H}_2\text{O})_6]^{2+}$  because the high spin complexes of  $\text{Co}^{2+}$  ( $d^7$ ) are labile and undergo rapid exchange with the solvent which is present in great excess.

In these cases, it is found that X is transferred quantitatively from the coordination sphere of  $\text{Co}^{3+}$  to that of  $\text{Cr}^{2+}$  as electron transfer is achieved. Therefore, it is likely that electron transfer occurs through a bridging ligand that is simultaneously part of the coordination spheres of both metal ions and that the bridging group remains as part of the coordination sphere of the inert complex of  $\text{Cr}^{3+}$  that is produced. In essence, the electron is thus “conducted” through that ligand.

In reactions such as that described above, rates of electron transfer are found to depend on the nature of X, and the rate varies in the order  $I^- > Br^- > Cl^- > F^-$ . However, other reactions are known where the opposite trend is observed. There are undoubtedly several factors involved, and these include  $F^-$  forming the strongest bridge but  $I^-$  being the best “conductor” for the electron being transferred because it is much easier to distort the electron cloud of  $I^-$  (it is much more polarizable and has a lower electron affinity). Therefore, in different reactions these effects may take on different weights leading to variations in the rates of electron transfer that do not follow a particular order with respect to the identity of the anion.

### 21.3 A CLOSER LOOK AT SUBSTITUTION REACTIONS

As has already been discussed, ligands function as Lewis bases when they donate pairs of electrons to a metal ion which behaves as a Lewis acid, and the coordinate bonds are the result of electron pair donation and acceptance. When  $NH_3$  is added to a solution containing  $[Cu(H_2O)_4]^{2+}$ , the color changes to a dark blue because of the formation of  $[Cu(NH_3)_4]^{2+}$  as a result of the *substitution* of  $NH_3$  molecules for  $H_2O$  molecules. In fact, the most frequently encountered type of reaction of complexes is the substitution of one ligand for another.

Lewis bases are called *nucleophiles* because of their attraction to a center of positive charge, and a positive species that attracts a pair of electrons is called an *electrophile* (see Chapter 6). Therefore, when one Lewis base replaces another, the reaction is called a *nucleophilic substitution*. In some cases, reactions of this type may be designated as  $S_N1$  or  $S_N2$  depending on whether the reaction is first order or second order, respectively. A nucleophilic substitution reaction involving a complex can be shown as



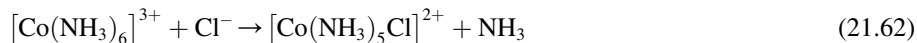
where X is the leaving ligand and Y is the entering ligand.

Substitution reactions are also known in which one metal displaces another as a result of the difference in their strengths as Lewis acids. A general reaction of this type can be represented by the equation

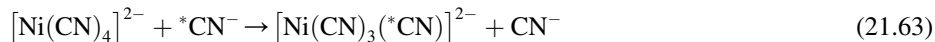


Because metal ions are electrophiles, a reaction of this type is known as *electrophilic substitution* (the mechanism may be described as  $S_{E1}$  or  $S_{E2}$ ). In this reaction, the ligands are transferred from one metal to another which is equivalent to one metal replacing the other. This type of reaction is much less common than the nucleophilic substitution reactions that will be considered in this section.

Substitution reactions take place with rates that differ enormously. For example, the reaction



is very slow whereas the exchange of cyanide ions (where  $*CN^-$  is an isotopically labeled cyanide ion) in the reaction



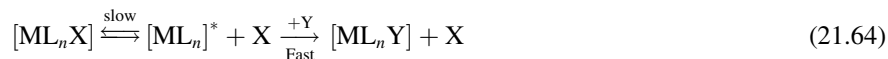
is very fast. Complexes that undergo rapid substitution reactions are said to be *labile* and those for which substitution reactions are slow are described as *inert*. The term “inert” does not mean that the complex is totally inert and that substitution reactions do not occur. It means that substitution in such complexes takes place *slowly*.

For a chemical reaction, there may also be competing effects with regard to kinetics and thermodynamics. Consider a reaction in which there are two possible products that have different energies. In general, the more stable product would be expected to be the main product of the reaction. However, the *rate* of a chemical reaction depends on the pathway. In order for reactants to be transformed into products, there is an energy barrier (the activation energy required to form the transition state) over which the reactants must pass. It is possible that in some cases, the rate is *higher* for the formation of a *less* stable product because of a lower activation energy for that reaction. In such cases, the less stable product is formed more rapidly so more of it is produced. The product obtained is referred to as the “kinetic product” and the more stable product is called the “thermodynamic product.” As a result, the products of the reaction may be the result of the kinetic factors rather than the thermodynamic stability of the product.

Substitution reactions may take place by mechanisms that represent two ideal or limiting types. In the first of these processes, the leaving group departs prior to the entering ligand becoming attached. The rate of such a substitution process is limited only by the concentration of the starting complex. In this mechanism, the coordination number of the metal is



reduced by one in forming the transition state. Such a mechanism is called a *dissociative pathway*, and the process can be illustrated by the equation



The transition state consists of the complex  $\text{ML}_n$  as a result of X having been lost. Because the overall rate of substitution is dependent on the concentration of the transition state, which depends only on the concentration of the starting complex, the reaction follows the rate law

$$\text{Rate} = k[\text{ML}_n\text{X}] \quad (21.65)$$

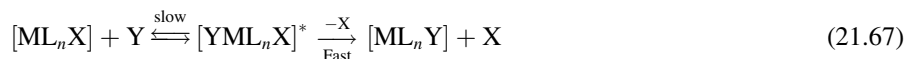
The reactant and the transition state represent two states that have different energy. Therefore, the population of the state of higher energy is determined by the Boltzmann Distribution Law, and the rate constant is expressed by the exponential equation known as the Arrhenius Equation,

$$k = Ae^{-E/RT} \quad (21.66)$$

where  $E$  is the activation energy,  $A$  is the frequency factor,  $R$  is the gas constant, and  $T$  is the temperature (K). The energy profile for this type of reaction is shown in [Figure 21.1](#).

A process following the rate law shown in [Eq. \(21.65\)](#) is said to be an  $\text{S}_{\text{N}}1$  (substitution, nucleophilic, unimolecular) process. The term *unimolecular* refers to the fact that a single species is required to form the transition state. Because the rate of such a reaction depends on the rate of dissociation of the M–X bond, the mechanism is also known as a *dissociative pathway*. In aqueous solutions, the solvent is also a potential nucleophile, and it solvates the transition state. In fact, the activated complex in such cases would be indistinguishable from the aqua complex  $[\text{ML}_n\text{H}_2\text{O}]$  in which a molecule of  $\text{H}_2\text{O}$  actually completes the coordination sphere of the metal ion after X leaves. This situation is represented by the dotted curve in [Figure 21.1](#) where the aqua complex is an *intermediate* that has lower energy than  $[\text{ML}_n]^*$ . The species  $[\text{ML}_n\text{H}_2\text{O}]$  is called an *intermediate* because it has a lower energy than that of the activated complex,  $[\text{ML}_n]$ .

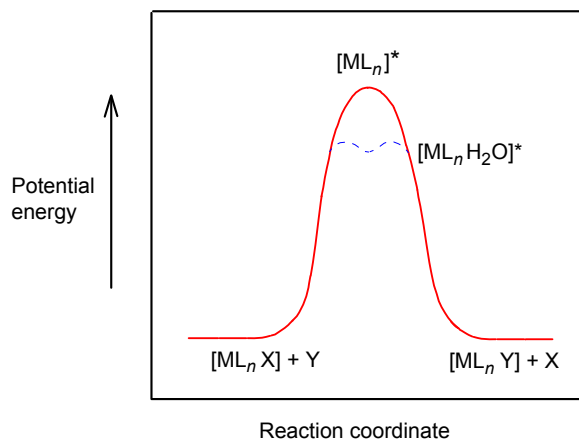
For numerous substitution processes, the reaction can be represented by the general equation



In this case, the formation of the activated complex is dependent on the concentrations of both  $\text{ML}_n\text{X}$  and Y giving rise to the rate law

$$\text{Rate} = k[\text{ML}_n\text{X}][\text{Y}] \quad (21.68)$$

Such a process is called an  $\text{S}_{\text{N}}2$  (substitution, nucleophilic, bimolecular) process. The term *bimolecular* indicates that two species are required to form the transition state. Because the transition state is formed by the “association” of the two reactant species, this type of reaction mechanism is also known as an *associative pathway*. In an  $\text{S}_{\text{N}}2$  process, the metal has a



**FIGURE 21.1** The energy profile for a dissociative ( $\text{S}_{\text{N}}1$ ) process.

higher coordination number in the transition state than it does in either the initial complex or product. The energy profile for this type of process is shown in Figure 21.2.

The fact that the solvent may participate in an  $S_N1$  process has already been mentioned. If that is the case, the overall process may be represented by the following equations.



The process actually involves two bimolecular processes each of which requires solvent participation. However, the first step may also be a first-order step in which  $\text{H}_2\text{O}$  coordinates only after  $\text{X}$  leaves. Although the reaction has been shown as if it takes place in water, other solvents could also be used, and a great difference in the overall rate of reaction should be expected. Because the coordinating abilities of solvents are different, the rates of the substitution reactions depend on the solvent chosen. Although many reactions are more complicated than suggested by the  $S_N1$  and  $S_N2$  models, there are many substitution reactions for which these rate laws (the so-called *limiting* cases) are applicable.

## 21.4 SUBSTITUTION IN SQUARE PLANAR COMPLEXES

Although in the previous section the basic concepts related to substitution reactions were explained with reference to octahedral complexes, substitution reactions are also common in square planar complexes. Studies on these complexes have resulted in a great deal of knowledge of the mechanisms of these reactions so a brief description of the topic will be presented next.

### 21.4.1 Mechanisms

In Chapter 20, it was explained that the most common complexes having square planar geometry are those of the  $d^8$  ions Pt(II), Pd(II), Au(III), and to a lesser extent, Ni(II). Substitution reactions of complexes containing Pt(II) have been the subject of a large number of kinetic studies.

Consider the substitution reaction represented by the equation



It can easily be seen that the only product possible when one X ligand is replaced is  $[\text{PtX}_3\text{Y}]$ . The first of the limiting mechanisms possible for such a substitution reaction is described as X leaving the coordination sphere of the metal ion before Y enters. This is the  $S_N1$  or dissociative mechanism described earlier, and it can be shown as

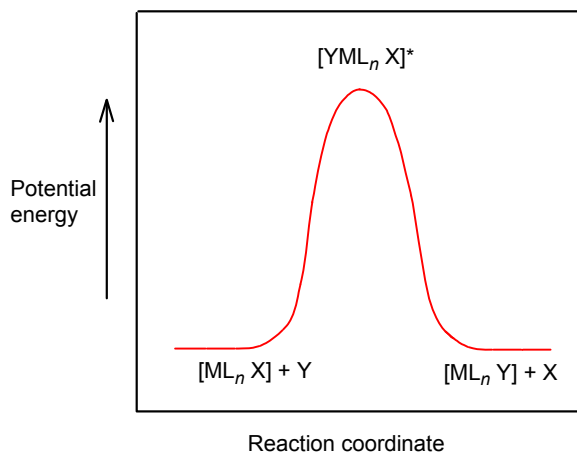
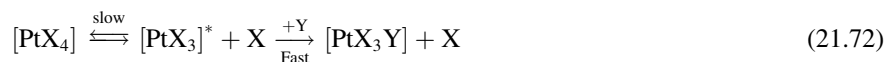
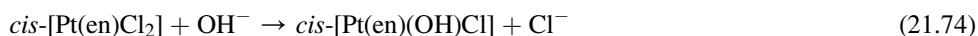


FIGURE 21.2 The energy profile for an associative ( $S_N2$ ) process.

The rate law corresponding to this mechanism is

$$\text{Rate} = k_1[\text{PtX}_4] \quad (21.73)$$

and it shows that the rate depends only on the concentration of  $\text{PtX}_4$ . Actually, this is true only when there is sufficient  $\text{Y}$  present to make the second step fast compared to the first, and that occurs when the concentration of  $\text{Y}$  is large compared to that of the transition state. The concentration of the transition state in a reaction is normally very low, but if the concentration of  $\text{Y}$  were also very low, the two concentrations might be of comparable magnitude. In such a case, the second step would not be much faster than the first. In that case, the reaction would not appear to be a first-order,  $\text{S}_{\text{N}}1$ , dissociative process but rather a second-order process. However, the reaction

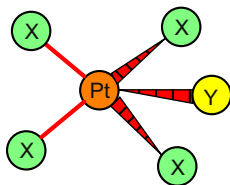


follows a rate law

$$\text{Rate} = k_1[\text{Pt}(\text{en})\text{Cl}_2] \quad (21.75)$$

so an  $\text{S}_{\text{N}}1$  mechanism is indicated.

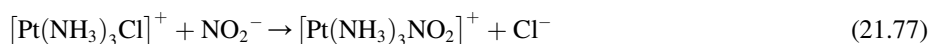
If  $\text{Y}$  substitutes for  $\text{X}$  in  $[\text{PtX}_4]$  by an  $\text{S}_{\text{N}}2$  mechanism, the process involves an increase in coordination number of the metal as the transition state forms. The entering group,  $\text{Y}$ , becomes coordinated before the leaving group,  $\text{X}$ , departs giving a transition state that can be shown as



In this associative pathway, the rate law depends on the concentrations of both the complex ion and the entering ligand because both are present in the transition state. Therefore, the rate law for the reaction is

$$\text{Rate} = k_2[\text{PtX}_4][\text{Y}] \quad (21.76)$$

For the reaction



the observed rate law is

$$\text{Rate} = k_2[\text{Pt}(\text{NH}_3)_3\text{Cl}^+][\text{NO}_2^-] \quad (21.78)$$

indicating that the process takes place by an  $\text{S}_{\text{N}}2$  or associative pathway. Dissociative processes frequently take place at rates that are essentially independent of the nature of the entering ligand because the rate is determined by the loss of the leaving ligand. Rates of reactions that take place by an associative mechanism depend on the nature of the entering ligand and its ability to bond to the metal as the transition state forms.

For most complexes of  $\text{Pt}(\text{II})$ , the rate of substitution varies with the nature of the entering group and the processes are  $\text{S}_{\text{N}}2$ . However, in some cases substitution reactions are found to follow a rate law of the form

$$\text{Rate} = k_1[\text{Complex}] + k_2[\text{Complex}][\text{Y}] \quad (21.79)$$

where  $\text{Y}$  is the entering ligand. The form of this rate law indicates the reaction is taking place simultaneously by a dissociative *and* an associative pathway. In reality, this may not be the case, and it is likely that the first term in the rate law represents a pseudo first-order step. The first term may actually represent a second-order process that is first order in complex and first order in solvent. The solvent can later be displaced by the entering ligand, so the actual rate law is more accurately represented as

$$\text{Rate} = k_1[\text{Complex}][\text{Solvent}] + k_2[\text{Complex}][\text{Y}] \quad (21.80)$$

With a rate law having this form, the pseudo first-order (with respect to the complex) rate constants can be written as

$$k = k_1 + k_2[\text{Y}] \quad (21.81)$$

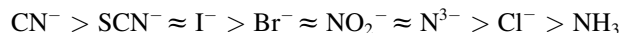
and the rate constant should show a marked dependence on the nature of Y for a series of reactions carried out at the same temperature and concentration when different entering ligands are used. In accord with this, the substitution reactions,



where pip is piperidine,  $\text{C}_5\text{H}_{10}\text{NH}$ , show a great variation depending on the nature of Y. The rate constants for a series of different entering ligands, Y, are shown as follows.

Y	$\text{Br}^-$	$\text{SCN}^-$	$\text{SeCN}^-$	$\text{SC}(\text{NH}_2)_2$
$k, \text{M}^{-1} \text{s}^{-1}$	0.0069	0.400	3.3	4.6

The data shown above are in accord with an associative mechanism of substitution with the ligands that bond preferentially to Pt(II) giving faster substitution reactions. From these and similar studies, it is possible to arrange a series of ligands on a scale of relative rates of substitution with the order being



In a general way, the rate of substitution in Pt(II) complexes increases with the softness of the entering ligand. Also, the ability of the ligand to form  $\pi$  bonds with the metal affects the rate of its entry into the coordination sphere of the metal.

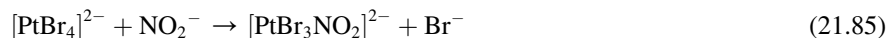
When a particular ligand is being replaced from complexes of the  $d^8$  ions Pt(II), Pd(II), or Ni(II), the replacement will be fastest for the Ni(II) complex and slowest for the Pt(II) complex. In fact, complexes of Pt(II) are among the most stable (and hence most kinetically inert) known. In most cases, forming a transition state by altering the structure of the starting complex is accompanied by a loss of ligand field stabilization energy (LFSE). Therefore, it would be expected that substitution reactions of complexes of Ni, Pd, and Pt would vary in the order  $\text{Ni} > \text{Pd} > \text{Pt}$  because the LFSE is far greater for second and third row metal ions than for first row metal ions (see Chapter 20).

### 21.4.2 The *Trans* Effect

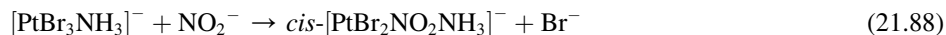
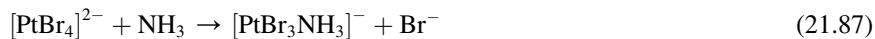
If a complex such as  $\text{PtX}_4$  has one of the ligands X replaced by Y, a single product,  $\text{PtX}_3\text{Y}$ , results. However, replacement of a second X by Y can lead to two different products.



Specific examples of this behavior are shown in the following equations.



If the order of introducing  $\text{NO}_2^-$  and  $\text{NH}_3$  into the  $[\text{PtBr}_4]^{2-}$  complex is reversed, it is found that the result is quite different. In that case, the reactions can be shown as follows.



Although the only difference in the two cases is the order in which the  $\text{NH}_3$  and  $\text{NO}_2^-$  ligands are added, this difference is significant enough to cause one of the products to have a *cis* configuration and the other to have a *trans* structure. It appears that if  $\text{NO}_2^-$  is introduced first, it causes the  $\text{Br}^-$  *trans* to it to be loosened enough so that when  $\text{NH}_3$  enters, it enters the coordination sphere *trans* to  $\text{NO}_2^-$ . In the second case, the complex contains  $\text{Br}^-$  and  $\text{NH}_3$  as ligands at the time  $\text{NO}_2^-$  enters the coordination sphere. In this case,  $\text{NO}_2^-$  enters *trans* to  $\text{Br}^-$  but *cis* to  $\text{NH}_3$ . Consequently, it appears that one  $\text{Br}^-$  causes the  $\text{Br}^-$  *trans* to it to be loosened more than  $\text{NH}_3$  causes the  $\text{Br}^-$  *trans* to it to be loosened. An alternative explanation could be that in the first case  $\text{Br}^-$  exerts a strong labilizing influence on the position *cis* to it so that  $\text{NH}_3$  enters *cis* to  $\text{Br}^-$ , but *trans* to  $\text{NO}_2^-$ . In the second case, if  $\text{NH}_3$  were to labilize the group *cis* to it,  $\text{NO}_2^-$  would enter there and give *cis*- $[\text{PtBr}_2\text{NO}_2\text{NH}_3]$ . However, as will be shown later, the dominant effect is the *trans* labilization by  $\text{NO}_2^-$  and  $\text{Br}^-$  in the two cases, and both ligands exert a stronger *trans* effect than  $\text{NH}_3$ . The *trans* effect or *trans* influence is of major importance in

reactions of square planar complexes. Studies have shown that the general order of *trans* effect produced by several common ligands is



Although the product obtained in a substitution may have either a *cis* or *trans* configuration, it is possible in some cases for an isomerization to have taken place *after* the substitution step. Isomerizations of this type are much more common for complexes of Ni(II) than for those of Pd(II) or Pt(II) owing to the larger LFSE for the second and third row transition metal ions. In fact, it is safe to assume that substitution occurs with retention of configuration in virtually all cases involving square planar complexes of Pt(II). Other manifestations of the *trans* influence will now be explored.

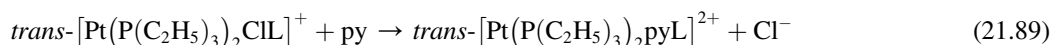
As described above, the *trans* effect is indicated by the fact that the products obtained when  $\text{NH}_3$  and  $\text{NO}_2^-$  are substituted for  $\text{Br}^-$  in  $[\text{PtBr}_4]^{2-}$  depend on the order of addition. If  $\text{NH}_3$  is added first followed by  $\text{NO}_2^-$ , the product is *cis*- $[\text{PtBr}_2\text{NH}_3\text{NO}_2]^-$ , but if  $\text{NO}_2^-$  is added first, the product is *trans*- $[\text{PtBr}_2\text{NH}_3\text{NO}_2]^-$ . Clearly the stereochemistry of the product is the result of a *trans* effect. If this *trans* directing influence is manifested in this way, it should also be evident from other properties of square planar complexes. In fact, there should be kinetic as well as thermodynamic and structural evidence to indicate the difference in *trans* influence exerted by different ligands. In the case of square planar complexes, much evidence of this type actually exists and some of it will be described here. For octahedral complexes, the situation is by no means clear as will also be described.

One of the definitive studies demonstrating the *trans* influence on metal–ligand bonds made use of infrared spectroscopy to study metal–ligand stretching frequencies. In this study, the positions of the stretching bands associated with the Pt–H bonds were determined for a series of complexes of the type *trans*- $[\text{Pt}(\text{P}(\text{C}_2\text{H}_5)_3)_2\text{HL}]$ , where L is one of a series of ligands in the position *trans* to the hydrogen atom. In these complexes, a *trans* influence caused by L should cause the Pt–H bond to be altered slightly, and the positions of the stretching bands reflect the difference as shown by the following data.

Group <i>trans</i> to H	$\text{Cl}^-$	$\text{Br}^-$	$\text{I}^-$	$\text{NO}_2^-$	$\text{SCN}^-$	$\text{CN}^-$
Pt–H stretching bond, $\text{cm}^{-1}$	2183	2178	2156	2150	2112	2042

The data provide a clear indication that the group *trans* to H does produce an effect on the Pt–H bond. In fact, the position of the Pt–H bond when  $\text{CN}^-$  is *trans* to it shows that the  $\text{CN}^-$  weakens the Pt–H bond considerably. On the other hand, the chloride ion, being a ligand that bonds less strongly than  $\text{CN}^-$ , changes the Pt–H bond *trans* to it very little. It follows that as a result of the cyanide ion weakening the Pt–H bond, the H would be more easily replaced when  $\text{CN}^-$  is *trans* to it than when  $\text{Cl}^-$  is in the *trans* position. The relative *trans* influence indicated by the change in the Pt–H bonds corresponds to the order obtained for the ligands from other types of data.

Studies on many replacement reactions using square planar complexes have been carried out. For the reaction



the observed rate law is

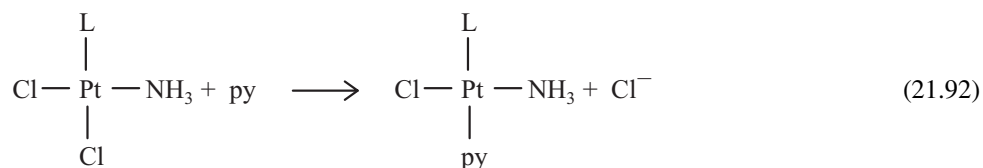
$$\text{Rate} = k_1[\text{Complex}] + k_2[\text{Complex}][\text{py}] \quad (21.90)$$

where the pseudo first-order rate constant is

$$k_{\text{obs}} = k_1 + k_2[\text{py}] \quad (21.91)$$

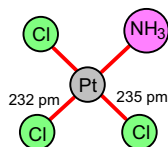
The values of  $k_1$  and  $k_2$  are found to vary greatly when the nature of L changes. For example, when  $\text{H}^-$  is *trans* to  $\text{Cl}^-$ ,  $k_1 = 1.8 \times 10^{-2} \text{ s}^{-1}$  and  $k_2 = 4.2 \text{ M}^{-1} \text{ s}^{-1}$ . When  $\text{L} = \text{Cl}^-$ ,  $k_1 = 1.0 \times 10^{-6} \text{ s}^{-1}$  and  $k_2 = 4.0 \times 10^{-4} \text{ M}^{-1} \text{ s}^{-1}$ . Thus, the rate constants vary by a factor of about  $10^4$  depending on the nature of the group *trans* to the leaving  $\text{Cl}^-$ .

The kinetic differences produced by a group in the *trans* position have also been studied for other processes. For example, for the reaction



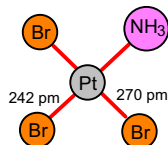
The activation energies were found to be 79, 71, and 46 kJ mol<sup>-1</sup> when L is Cl<sup>-</sup>, Br<sup>-</sup>, or NO<sub>2</sub><sup>-</sup>, respectively. The rate constants for the reactions were found to be 6.3, 18, and 56 M<sup>-1</sup> s<sup>-1</sup> when L is Cl<sup>-</sup>, Br<sup>-</sup>, or NO<sub>2</sub><sup>-</sup>, respectively.

In addition to the effects on reaction rates, structural differences occur as a result of the *trans* effect. In [PtNH<sub>3</sub>Cl<sub>3</sub>]<sup>-</sup> the Pt–Cl bond distances are as given below.



The fact that the Pt–Cl bond *trans* to NH<sub>3</sub> is shorter than that opposite a Pt–Cl bond shows that NH<sub>3</sub> exerts a smaller *trans* effect than does Cl<sup>-</sup>.

For the complex [PtNH<sub>3</sub>Br<sub>3</sub>]<sup>-</sup>, the Pt–Br distances are as shown in the structure.



From the bond lengths, it is clear that Br<sup>-</sup> weakens the Pt–Br bond *trans* to it much more than the NH<sub>3</sub> does the Pt–Br bond *trans* to it. The difference in the bond lengths in the bromide complex (28 pm) is greater than in the chloride complex (3 pm) indicating that Br<sup>-</sup> exerts a stronger *trans* influence than does Cl<sup>-</sup>.

The evidence cited previously shows clearly that there is a *trans* influence in square planar complexes and that the magnitude depends on the nature of the ligand. In fact, the *trans* directing influence is a factor of major importance in the structures and substitution reactions of square planar complexes. It is now necessary to provide an explanation of how this phenomenon is caused and how it is related to the nature of the ligands.

### 21.4.3 Causes of the *Trans* Effect

There are two lines of reasoning that have been utilized to explain the *trans* effect. The first of these arguments deals primarily with the charge distribution and polarizability of the ligand causing the *trans* effect. According to this theory, the ligand Y in *trans*-[PtA<sub>2</sub>XY] produces a *trans* effect by weakening the Pt–X bond when Y has a greater polarizability than X. This may be illustrated as shown in Figure 21.3.

In this case, the metal ion carries an overall positive charge (+2 in the case of the Pt complexes described earlier). If Y is easily polarizable, the +2 charge of the metal ion will cause an unsymmetrical charge distribution in the ligand Y. However, because Pt<sup>2+</sup> is a large, polarizable ion, there will be a migration of the electrons in Pt<sup>2+</sup> away from the ligand Y as the electron density in the ligand shifts toward Pt<sup>2+</sup>. Because X is *less* polarizable than Y, there will be less shifting of electron density in X and the result is that X now faces a side of the metal ion that is slightly less positive, which weakens the Pt–X bond. As a result, Y causes X to be less strongly bonded and more easily removed from the coordination sphere. Because the primary influence of Y is the weakening of the Pt–X bond,

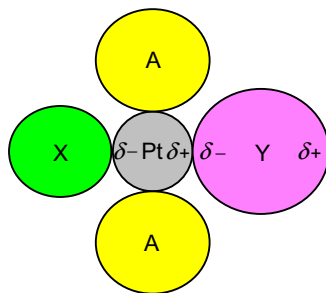


FIGURE 21.3 Charge distribution in the X–Pt–Y linkage.

the polarization theory is sometimes referred to as the *theory of bond weakening*. Some of the effects of this bond weakening have already been seen from the data presented earlier. It has been shown that for the halide ions, the order of decreasing *trans* effect is  $I^- > Br^- > Cl^-$ , and this order is in accord with the decreasing polarizability of these ions. The  $H^-$  ion is extremely polarizable (it has *twice* as many electrons as protons and is about the size of  $I^-$ ) so it exerts a very strong *trans* effect. However,  $CN^-$  exerts a very strong *trans* effect, but its polarization is only moderate. Therefore, the polarization theory alone is too simple to deal with the effect produced by all types of ligands.

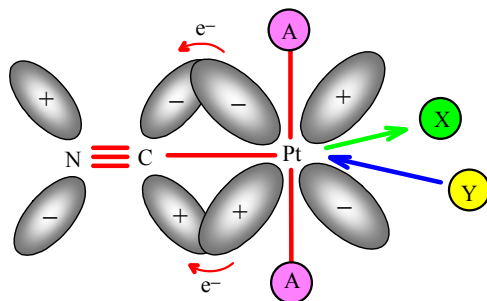
A different explanation for the *trans* effect is based on the ability of a ligand to form multiple bonds to the metal ion. In Chapter 20, multiple bonding of such ligands as CO and  $CN^-$  to metal ions by back donation of electron density from the *d* orbitals of the metal to vacant  $\pi^*$  orbitals on the ligands was described. The removal of electron density from the metal ion by a ligand should not significantly weaken bonds to the other ligands except for the one *trans* to the  $\pi$ -bonding ligand. However, the ligand in the *trans* position competes for the electron density from the *d* orbitals on the metal. It should be remembered that the true criterion for a *trans* effect is that one ligand is easier to remove from the complex because of the *trans* influence of another. Accordingly, the explanation of the strong *trans* effect lies in the stabilization of the transition state for the substitution reaction. This lowering of the activation energy causes a replacement reaction to proceed faster under a given set of conditions when the ligand *trans* to the leaving group has greater  $\pi$ -bonding ability. The stabilization of the transition state by  $\pi$  bonding is shown in Figure 21.4.

The removal of electron density from the Pt(II) by multiple bonding to CN causes the side of the Pt(II) toward X to be electron deficient and more susceptible to nucleophilic attack by strong nucleophiles. Thus, the activation energy for forming the transition state is lower and the reaction proceeds faster. For a large number of ligands the apparent order of *trans* effect produced is



The overall *trans* effect produced by a ligand is the result of both its polarizability and its  $\pi$  bonding ability. In the case of  $H^-$ , which cannot form  $\pi$  bonds, the large *trans* effect is the result of it being very polarizable. For CO and  $CN^-$ , the availability of  $\pi^*$  orbitals for forming  $\pi$  bonds to the metal causes the large *trans* effect. Thus, the overall ability of a ligand to cause a *trans* effect depends on both of these factors working together.

It should also be pointed out that the *trans* effect is not limited to complexes of Pt(II) but should exist in any square planar complex. Generally, the *trans* effect is smaller in complexes of Pd(II) than in those of Pt(II) and this is exactly as would be expected from the fact that Pd(II) is less polarizable (harder) than Pt(II). In terms of the  $\pi$  bonding hypothesis, the Pd(II) is a harder acid than Pt(II) (and less likely to participate in back donation) and does not form covalent bonds as well. By either way of explaining the *trans* effect, the effect should be smaller in complexes of Pd(II) than in those of Pt(II). The *trans* effect has also been investigated for square planar complexes of the  $d^8$  ion Au(III) and, in general, it appears to be similar to that in Pd(II). Because Au(III) has a higher positive charge than Pt(II), it should be harder and the lower polarizability should result in a smaller *trans* effect than in Pt(II) complexes. This is observed and the *trans* effect for Au(III) appears to be similar to that found in complexes of Pd(II) although palladium is a second row transition metal whereas gold is a third row metal.



**FIGURE 21.4** Stabilization of the transition state by  $\pi$  bonding by a ligand (in this illustration it is  $CN^-$ ) *trans* to the leaving group. In this case, the shift of electron density away from X makes it easier to remove.

## 21.5 SUBSTITUTION IN OCTAHEDRAL COMPLEXES

### 21.5.1 Classification Based on Rates

In a general way, the ease of substitution in a metal complex is related to the electronic structure of the metal ion. Labile complexes are usually those having  $sp^3d^2$  (outer orbital) hybrid bond types or those having  $d^2sp^3$  (inner orbital) hybrid types with one or more of the inner  $d$  orbitals being vacant. In these cases, an orbital having a similar energy to the valence orbitals that can form a bond to the entering ligand is available. Thus, inner orbital complexes having  $d^0$ ,  $d^1$ , and  $d^2$  metal ion configurations should undergo rapid substitutions by an associative process. In agreement with this line of reasoning, complexes of  $Sc^{3+}$  ( $d^0$ ),  $Ti^{3+}$  ( $d^1$ ), and  $V^{4+}$  ( $d^2$ ) are *labile*.

Low spin complexes in which the metal configuration is  $d^3$ ,  $d^4$ ,  $d^5$ , or  $d^6$  would require either a ligand to leave before substitution could occur or the utilization of an outer  $d$  orbital to form a seven-bonded transition state. Such complexes frequently undergo substitution by a dissociative process because expanding the coordination number of the metal is made difficult by the lack of a vacant orbital of suitable energy. In either the  $S_N1$  or  $S_N2$  case, the substitution should be much slower than it is for labile complexes. Accordingly, complexes of  $V^{2+}$ ,  $Cr^{3+}$ ,  $Mn^{4+}$  (all of which are  $d^3$  ions), low spin complexes of  $Co^{3+}$ ,  $Fe^{2+}$ ,  $Ru^{2+}$ ,  $Rh^{3+}$ ,  $Ir^{3+}$ ,  $Pd^{4+}$ , and  $Pt^{4+}$  (all of which are  $d^6$  ions) and low spin complexes of  $Mn^{3+}$ ,  $Re^{3+}$ , and  $Ru^{4+}$  are classified as *inert*. Inert does not mean that substitution does not occur, but rather that it occurs much more slowly than it does in labile complexes.

The rate constants for substitution reactions of labile complexes are frequently in the order of  $10^1$ – $10^6$   $M^{-1} s^{-1}$ , whereas those for inert complexes are as low as  $10^{-5}$ – $10^{-8}$   $s^{-1}$ . Certainly the difference in electronic structures is one factor contributing to this enormous variation in rates of substitution. Other reasons will be discussed later in this chapter.

Gray and Langford (1968) have classified metal ions into four categories based on the rates of  $H_2O$  exchange between their aqua complexes and bulk solvent. The classifications that they have proposed are shown in Table 21.1. Exchange reactions of the first three groups of metal ions are so rapid that they cannot be studied by the ordinary types of experimental techniques used to follow reaction kinetics. Continuous flow, stopped flow, and NMR line-broadening techniques have been most useful for studying these fast reactions. Complexes of the metals in the first two groups would be appropriately considered as labile. Because the interaction between the metals and ligands in the first three groups is predominantly electrostatic in nature, there is a general decrease in the rate of exchange as the charge to size ratio of the metal ion increases.

Exchange rates, and hence substitution reactions, with Class IV metals are slow enough to allow rate studies to be carried out by conventional kinetic techniques. This group of metal complexes has received an enormous amount of study, so it is natural that mechanistic information is available about these reactions. Consequently, the discussion of substitution mechanisms that follows is largely concerned with those complexes that are classified as inert or as belonging to the Group IV metal ions shown in Table 21.1.

### 21.5.2 The Effect of LFSE on Rate of Substitution

Some of the manifestations of the LFSE have already been described. It can be shown that producing a five-bonded transition state in a dissociative process or a seven-bonded transition state in an associative process would invariably lead to a loss of LFSE except for a few cases such as  $d^0$  or  $d^5$  high spin ions where the LFSE is zero. However, the loss of LFSE is different in a dissociative process depending on whether the five-bonded transition state is a trigonal bipyramid or a square-based pyramid (sometimes called a tetragonal pyramid).

**TABLE 21.1** Classification of Complexes Based on Rates of  $H_2O$  Exchange with Solvent

Class	Typical Range of First Order Rate Constant ( $s^{-1}$ )	Examples
I	$10^8$	$Na^+$ , $K^+$ , $Li^+$ , $Ba^{2+}$ , $Ca^{2+}$ , $Cd^{2+}$ , $Cu^+$
II	$10^4$ – $10^8$	$Mg^{2+}$ , $Mn^{2+}$ , $Fe^{2+}$ , $Zn^{2+}$
III	$10^0$ – $10^{-4}$	$Fe^{3+}$ , $Be^{2+}$ , $Al^{3+}$
IV	$10^{-1}$ – $10^{-9}$	$Co^{3+}$ , $Cr^{3+}$ , $Pt^{2+}$ , $Rh^{3+}$ , $Ir^{3+}$

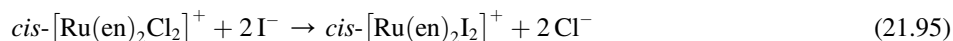
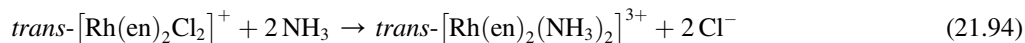
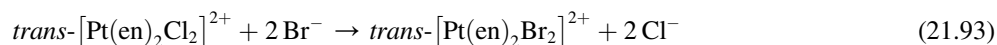


A complex containing a  $d^6$  ion surrounded by six ligands arranged octahedrally has an LFSE of  $-24 Dq$ , the largest LFSE observed. It can be shown that if one ligand is lost to form a trigonal bipyramid transition state in a dissociative process, the LFSE for the complex containing five ligands is only  $-12.52 Dq$ , and  $11.48 Dq$  of stabilization is lost. However, in forming a square-based pyramid transition state, only  $4 Dq$  is lost. Accordingly, *low spin* complexes of  $d^6$  ions such as  $\text{Co}^{3+}$  or  $\text{Fe}^{2+}$  would suffer a greater loss in LFSE than those having other electron configurations as the transition state containing five ligands is formed. Therefore, rates of substitution should be lower for complexes of  $\text{Co}^{3+}$  than they are for other types of complexes in which the LFSE is lower. This is in general agreement with experimental evidence regarding rates. It would also appear that the preferred transition state in the dissociative process is the square-based pyramid because only  $4 Dq$  of LFSE is lost in forming that structure from an octahedron. However, this is not always the case.

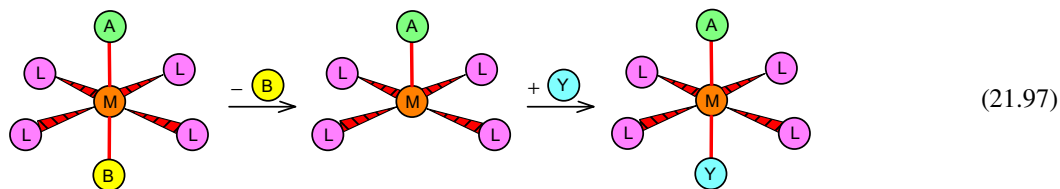
Because the LFSE is determined by the magnitude of  $Dq$ , differences in rates of substitution should be expected based on the nature of the ligands present. In general, this is observed. Also, because  $Dq$  increases from first to second to third row metals having the same electron configuration, it would be expected that the difference would result in substitution reactions being slower for complexes of those metals than those of metals in the first transition series. This is also a generally observed trend. One example of this is in the square planar complexes of  $\text{Ni}^{2+}$ ,  $\text{Pd}^{2+}$ , and  $\text{Pt}^{2+}$ , all  $d^8$  ions. Here  $Dq$  is greatest for  $\text{Pt}^{2+}$  and lowest for  $\text{Ni}^{2+}$  when the same ligands are present in each case. Thus, forming the transition state would result in the greatest loss in LFSE for  $\text{Pt}^{2+}$ . Although substitution reactions of complexes of these metals usually proceed by an associative mechanism, the rates decrease in the order  $\text{Ni}^{2+} > \text{Pd}^{2+} > \text{Pt}^{2+}$ , in accord with the variation in LFSE lost in forming the transition states. Although the same number of  $Dq$  units are lost in the case of each of the  $d^8$  metal ions, the magnitude of  $Dq$  is different for the various metals.

The loss of LFSE is only one of several energy factors involved in forming the transition state. Other factors may be of greater significance for specific complexes, and it would be misleading to imply that it is always the dominant factor. However, a large amount of evidence indicates that the loss of LFSE in forming the transition state is a dominant factor in many substitution reactions.

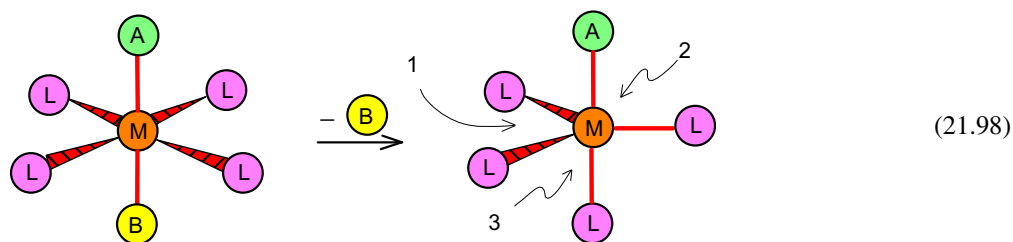
One strong indication of the importance of ligand field effects is the fact that substitution reactions in octahedral complexes of  $\text{Pt}^{4+}$ ,  $\text{Rh}^{3+}$ , and  $\text{Ru}^{3+}$  occur without rearrangement. The following are examples of reactions of this type.



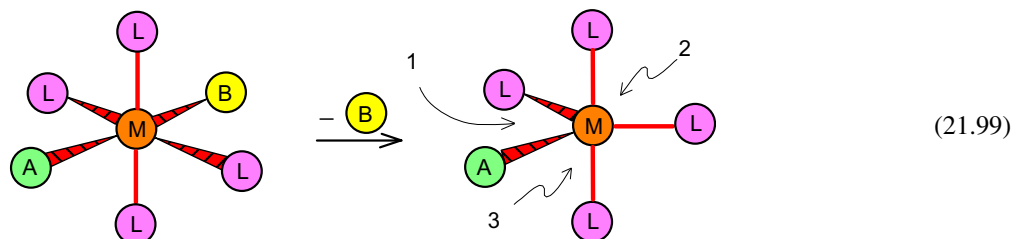
These reactions show that there is no loss of configuration as substitution occurs. For these second and third row metals, splitting of the  $d$  orbitals produced by en and  $\text{Cl}^-$  is considerably larger than it is in the case of first row metals. As mentioned previously, the formation of a square-based pyramid transition state is accompanied by a smaller loss in LFSE than is the formation of a trigonal bipyramid transition state. Thus, attack by the entering ligand on a square-based pyramid transition state produces a product having the same configuration as the starting complex as shown in Eq. (21.97).



In the trigonal bipyramid transition state, the attack can be along any side of the trigonal plane (where there is usually more free space), so it can result in either a *cis* or *trans* product. This situation is shown as follows where the ligand A is in an *axial* position in the transition state.



Attack at any of the positions in the equatorial plane labeled 1, 2, or 3 produces a product that has the entering ligand in a position *cis* to A. Therefore, the product when the entering ligand is X would be expected to be exclusively *cis*-[ML<sub>4</sub>AX]. However, if A is in an *equatorial* position in the *transition state*, the situation is different because the three positions of attack by the entering ligand are not all equivalent.

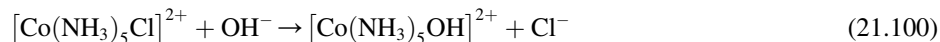


Attack at position 2 in the equatorial plane as shown above produces a product in which the entering ligand will be *trans* to A, but attack at position 1 or 3 would lead to a *cis* product. Therefore, the product when X is the entering ligand should be 2/3 *cis*- and 1/3 *trans*-[ML<sub>4</sub>AX]. It should be mentioned that if the ligands in the trigonal plane are very large, steric factors can make it possible that entry of a ligand is not equally probable in all three positions. Therefore, the product distribution may not be what it is expected to be according to the number and type of sites available in the transition state.

It appears that the square-based pyramid transition state is consistent with the fact that substitution in these second and third row metal complexes occurs without isomerization. This is primarily due to the fact that less LFSE is lost in forming a transition state when this structure is formed. On the other hand, Co<sup>3+</sup> and Cr<sup>3+</sup> being first row metals may have the same *number* of *Dq* units lost in forming the transition state as do Ir<sup>3+</sup> or W<sup>3+</sup>, but *Dq* is much smaller for the first row metal. Thus, the trigonal bipyramid structure *may* be formed more easily for first row metals, and if it is, the substitution can lead to either a *cis* or *trans* product. However, the interaction of the solvent with the transition state may offset part of the loss of LFSE. Accordingly, substitution reactions of complexes containing first row metals are frequently accompanied by a change in structure.

### 21.5.3 The S<sub>N</sub>1CB Mechanism

For low spin complexes of Co<sup>3+</sup> (*d*<sup>6</sup>), many substitution reactions occur at a rate that does not depend on the nature of the entering ligand, and the slow reactions follow an S<sub>N</sub>1 rate law. However, the reaction of [Co(NH<sub>3</sub>)<sub>5</sub>Cl]<sup>2+</sup> with OH<sup>-</sup> in aqueous solutions is much faster than expected, and it can be shown as



Although this equation *appears* to represent the substitution of OH<sup>-</sup> for Cl<sup>-</sup>, the reaction follows the rate law

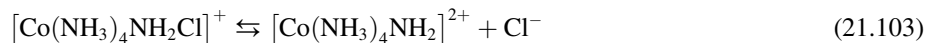
$$\text{Rate} = k[\text{Complex}][\text{OH}^-] \quad (21.101)$$

It thus appears that substitution of OH<sup>-</sup> in the cobalt complex follows a different mechanism than the dissociative pathway usually observed for substitution reactions of such complexes. Therefore, either OH<sup>-</sup> behaves differently than other nucleophiles or another way of explaining the observed rate law must be sought.

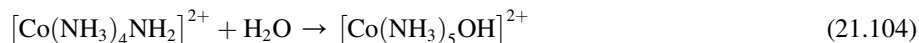
A novel mechanism that agrees with the observed rate law has been proposed for this reaction. In this mechanism, it is supposed that a rapid *equilibrium* is established in which a proton is removed from a coordinated NH<sub>3</sub> molecule to leave a coordinated NH<sub>2</sub><sup>-</sup>.



Consequently, this reaction produces the *conjugate base* (CB) of the starting complex. The dissociation of  $\text{Cl}^-$  from  $[\text{Co}(\text{NH}_3)_4\text{NH}_2\text{Cl}]^+$  occurs in an  $\text{S}_{\text{N}}1$  process.

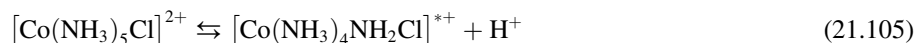


The reaction of this transition state with water occurs to complete the process by adding  $\text{H}^+$  to the coordinated  $\text{NH}_2^-$  ion and the attachment of  $\text{OH}^-$  to the metal ion.



The rate-determining step is that shown in Eq. (21.103), but the concentration of the CB is determined by the equilibrium shown in Eq. (21.102). Therefore, the concentration of the CB depends on the  $\text{OH}^-$  concentration, and that is indicated in the rate law.

For  $[\text{Co}(\text{NH}_3)_5\text{Cl}]^{2+}$  reacting as an acid (as  $\text{H}^+$  is removed), the equation showing the dissociation can be written as



and the equilibrium constant for the process can be represented as

$$K_a = \frac{[\text{CB}][\text{H}^+]}{[\text{Complex}]} \quad (21.106)$$

where CB is  $[\text{Co}(\text{NH}_3)_4\text{NH}_2\text{Cl}]^+$ , the CB of the starting complex. Solving this expression for [CB] gives

$$[\text{CB}] = \frac{K_a[\text{Complex}]}{[\text{H}^+]} \quad (21.107)$$

For an aqueous solution,  $[\text{H}^+][\text{OH}^-] = K_w$ , so  $[\text{H}^+]$  can be expressed as

$$[\text{H}^+] = \frac{K_w}{[\text{OH}^-]} \quad (21.108)$$

Substituting this expression for  $[\text{H}^+]$  in Eq. (21.107) gives

$$[\text{CB}] = \frac{K_a}{K_w}[\text{Complex}][\text{OH}^-] \quad (21.109)$$

Therefore, the rate law can be written as

$$\text{Rate} = k'[\text{CB}] = k' \frac{K_a}{K_w}[\text{Complex}][\text{OH}^-] = k[\text{Complex}][\text{OH}^-] \quad (21.110)$$

where  $k = k'K_a/K_w$ . Thus, the first-order dependence on  $\text{OH}^-$  is demonstrated although the actual substitution is still considered as  $\text{S}_{\text{N}}1$  or dissociative with respect to the cobalt complex. This mechanism is called the  $\text{S}_{\text{N}}1$  conjugate base ( $\text{S}_{\text{N}}1\text{CB}$ ) mechanism to denote that it involves the CB of the starting complex.

Experimental support for such a process is found in the fact that hydrogen atoms of the coordinated ammonia molecules are acidic enough to undergo some deuterium exchange in  $\text{D}_2\text{O}$ . Therefore, it is not unreasonable to expect a small concentration of the CB to be present in basic solutions. Also, if ligands other than  $\text{NH}_3$  are present and they contain no protons that can be removed, the rate of substitution should show no increase as  $[\text{OH}^-]$  increases. This is indeed the case for complexes such as *trans*- $[\text{Co}(\text{dipy})_2(\text{NO}_2)_2]^+$  in which the protons on the ligands are not subject to removal by a base (the hydrogen atoms attached to the rings are not removed). Other complexes have been studied, and it was found that the rate is enhanced by increasing  $[\text{OH}^-]$  only when the ligands contained hydrogen atoms that are susceptible to removal by a base. As a result of these and other studies, the  $\text{S}_{\text{N}}1\text{CB}$  mechanism has been shown to be applicable to substitution reactions of complexes in basic solutions when the complex contains ligands that have hydrogen atoms that can be removed.

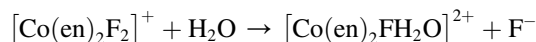
The material presented in this chapter provides an introduction to the vast area of reactions of coordination compounds. In addition to the types of reactions described, there is an extensive chemistry of reactions of coordinated ligands. Because many ligands are organic molecules, it is possible to carry out reactions on the ligands without disruption of the complexes, and some such reactions have been mentioned in this chapter. Several others will be shown in Chapter 22. Many reactions of coordination compounds have been studied in detail, and much is known about processes taking place in both solids and solutions. However, it is not possible in a book such as this to do more than introduce the field, but the listed references provide a basis for further study of this area.

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## PROBLEMS

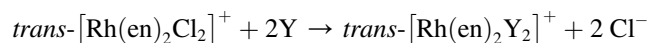
- Identify the type of reaction for each of the following processes.
  - $(\text{CO})_5\text{MnCH}_3 \rightarrow (\text{CO})_4\text{MnCOCH}_3$
  - $\text{HIr}(\text{CO})\text{Cl}(\text{PEt}_3)_2(\text{C}_2\text{H}_4) \rightarrow \text{Ir}(\text{CO})\text{Cl}(\text{PEt}_3)_2\text{CH}_2\text{CH}_3$
  - $\text{H}(\text{CO})_3\text{CoH}_2\text{C}=\text{CHR} \rightarrow (\text{CO})_3\text{CoCH}_2\text{CH}_2\text{R}$
- For each of the following reactions, draw the structure of the product that results.
  - trans*- $[\text{Ir}(\text{CO})\text{Cl}(\text{PEt}_3)_2]$  reacting with 1 mol of gaseous HBr
  - $\text{Mo}(\text{CO})_6$  reacting with excess pyridine,  $\text{C}_5\text{H}_5\text{N}$
  - The reaction of cobalt with CO at elevated temperature and pressure
- For each of the following reactions, draw the structure of the product that results.
  - The insertion reaction of  $\text{Fe}(\text{NO})(\text{CO})_3\text{Cl}$  with  $\text{SO}_2$
  - The reaction of  $\text{Co}(\text{CN})_5^{3-}$  with  $\text{H}_2\text{O}_2$
  - $\text{PtCl}_4^{2-}$  reacting with 1 mol of  $\text{NH}_3$  followed by 1 mol of  $\text{NO}_2^-$
- In contrast to other tetrahedral complexes, those of  $\text{Be}^{2+}$  undergo slow substitution, usually by an  $\text{S}_{\text{N}}1$  process. What is the basis for this behavior?
- Suppose a series of complexes of *trans*- $[\text{Pt}(\text{NH}_3)_2\text{LCl}]$  is prepared where L is  $\text{NH}_3$ ,  $\text{Cl}^-$ ,  $\text{NO}_2^-$ ,  $\text{Br}^-$ , or pyridine ( $\text{C}_5\text{H}_5\text{N}$ ). If the position of the Pt–Cl stretching bond is determined for each complex, arrange the values in the order of increasing wave number ( $\text{cm}^{-1}$ ) for the series of ligands L and explain your answer.
- The reaction



takes place much faster at pH = 2 than at pH = 4, although the reaction takes place slowly even in neutral solutions. Postulate a mechanism for this process and give the expected rate law. Discuss the difference between the acid catalysis observed here and the  $\text{S}_{\text{N}}1\text{CB}$  mechanism.

- Predict the configuration of the product for each of the following reactions.
  - $[\text{Pt}(\text{H}_2\text{O})_3\text{NO}_2]^+ + \text{NH}_3 \rightarrow$
  - $[\text{PtCl}_3\text{Br}]^{2-} + \text{NH}_3 \rightarrow$
  - $[\text{Pt}(\text{NH}_3)_3\text{Cl}]^+ + \text{CN}^- \rightarrow$

8. In *trans*-[Pd((CH<sub>3</sub>)<sub>3</sub>As)<sub>2</sub>Cl<sub>2</sub>] the Pd–Cl stretching bond is at 375 cm<sup>-1</sup> whereas for the corresponding *cis* complex it is at 314 cm<sup>-1</sup>. Explain this difference.
9. The reaction of 2 mol of (C<sub>2</sub>H<sub>5</sub>)<sub>3</sub>P with K<sub>2</sub>[PtCl<sub>4</sub>] produces a product having a different structure than is produced when 2 mol of (C<sub>2</sub>H<sub>5</sub>)<sub>3</sub>N react with K<sub>2</sub>[PtCl<sub>4</sub>]. Predict the structure in each case and explain the difference.
10. For the reaction

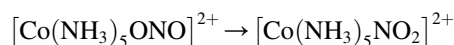


the following data were obtained.

Y (at 0.1 M)	I <sup>-</sup>	OH <sup>-</sup>	Cl <sup>-</sup>
10 <sup>5</sup> k (s <sup>-1</sup> )	5.2	5.1	4.0

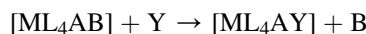
What mechanism does the reaction follow? Describe the transition state for this reaction. Would the same results be observed if Cr<sup>3+</sup> were the metal ion? Why or why not?

11. When the reaction



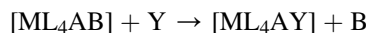
is studied under high pressure, it is found that the rate increases as the pressure increases. This indicates that the transition state occupies a smaller volume than the initial complex. Describe a mechanism that is consistent with that observation.

12. If the reaction



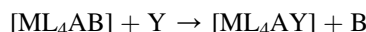
takes place with the formation of a trigonal bipyramid transition state with A in an axial position, what should be the distribution of *cis*-[ML<sub>4</sub>AY] and *trans*-[ML<sub>4</sub>AY] in the products? Provide a brief explanation for your answer.

13. If the reaction



takes place with the formation of a trigonal bipyramid transition state with A in an equatorial position, what should be the distribution of *cis*-[ML<sub>4</sub>AY] and *trans*-[ML<sub>4</sub>AY] in the products? Provide a brief explanation for your answer.

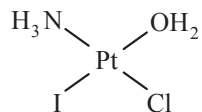
14. Suppose the reaction



in which A is a very large ligand (such as As(C<sub>6</sub>H<sub>5</sub>)<sub>3</sub>) takes place with the formation of a trigonal bipyramid transition state with A in an equatorial position, what would you expect the distribution of *cis*-[ML<sub>4</sub>AY] and *trans*-[ML<sub>4</sub>AY] in the products to be? Explain your answer.

15. If *cis*-[Pt(NH<sub>3</sub>)<sub>2</sub>ClBr] undergoes a substitution reaction with one SCN<sup>-</sup> replacing one NH<sub>3</sub>, draw the structure for the product. Why does the product not contain a Pt–NCS bond?
16. The synthesis of complexes containing the acetylacetonate ligand can be carried out in basic solutions. How does the base function? If the base is NH<sub>3</sub>, why is the product not a complex containing NH<sub>3</sub>?
17. Predict the product obtained by heating [Co(en)<sub>2</sub>ClH<sub>2</sub>O]SCN as a solid complex. How might the product obtained by heating [Pt(en)<sub>2</sub>ClH<sub>2</sub>O]SCN be different?
18. In this chapter, a series of ligands were arranged in order of their ability to produce a *trans* effect in complexes of Pt<sup>2+</sup>. Where would you place SeCN<sup>-</sup> in that series? Why?
19. When [Pt(NH<sub>3</sub>)Cl<sub>3</sub>]<sup>-</sup> undergoes two successive substitution reactions involving I<sup>-</sup> and H<sub>2</sub>O, what will be the product? Explain your answer.

20. Why is the *trans* effect less important for square planar complexes of  $\text{Ni}^{2+}$  than it is for complexes of  $\text{Pt}^{2+}$ ?
21. Draw the structure of the product obtained when *cis*- $[\text{Pt}(\text{NH}_3)_2\text{Cl}_2]$  undergoes an oxad reaction with  $\text{Cl}_2$ .
22. Starting with *cis*- $[\text{Pt}(\text{NH}_3)_2\text{Cl}_2]$ , devise a synthesis to prepare



23. Although  $\text{H}_2\text{C}=\text{CH}_2$  was not listed in the series of ligands causing a *trans* effect, explain where would you believe it fits.
24. For each of the following reactions, consider the ligand L to be a nonreacting molecule such as  $(\text{C}_6\text{H}_5)_3\text{P}$ . Draw the structure for the product in each case.
- $[\text{RhL}_3\text{Cl}] + \text{H}_2 \rightarrow$
  - $[\text{RhL}_3\text{Cl}] + \text{HCl}(g) \rightarrow$
  - $[\text{L}_2(\text{CO})\text{RhCH}_2\text{CH}_3\text{CO}] + \text{CO} \rightarrow$
  - $[\text{L}_2\text{Rh}(\text{CO})\text{H}] + \text{H}_2\text{C}=\text{CH}_2 \rightarrow$
25. Tell what type of reaction has occurred in each of the following cases and draw the structures of the products.
- $[\text{L}_2\text{Rh}(\text{CO})_2(\text{CH}_2\text{CH}_3)] \rightarrow [\text{L}_2\text{Rh}(\text{CO})(\text{COCH}_2\text{CH}_3)]$
  - $[\text{L}_2\text{Rh}(\text{CO})(\text{COCH}_2\text{CH}_3)] + \text{H}_2(g) \rightarrow [\text{L}_2\text{HRh}(\text{CO})(\text{COCH}_2\text{CH}_3)]$
  - $[\text{L}_2\text{Rh}(\text{H})(\text{CO})(\text{C}_2\text{H}_4)] \rightarrow [\text{L}_2\text{Rh}(\text{CO})(\text{CH}_2\text{CH}_3)]$
26. Tell whether the following complexes would be labile or inert and explain your answer in each case.
- $[\text{Fe}(\text{CN})_6]^{4-}$
  - $[\text{Cr}(\text{H}_2\text{O})_6]^{2+}$
  - $[\text{Ni}(\text{NH}_3)_6]^{2+}$
  - $[\text{CoF}_6]^{3-}$
  - $[\text{PtBr}_6]^{2-}$
  - $[\text{TiCl}_6]^{3-}$
27. The ligand  $\text{Et}_4\text{dien}$  is  $(\text{C}_2\text{H}_5)_2\text{NCH}_2\text{CH}_2\text{NHCH}_2\text{CH}_2\text{N}(\text{C}_2\text{H}_5)_2$ , and it forms stable chelates by coordinating through the three nitrogen atoms.
- Draw the structure for  $[\text{Pd}(\text{Et}_4\text{dien})\text{Cl}]^+$ .
  - The hydrolysis reaction of  $[\text{Pd}(\text{Et}_4\text{dien})\text{Cl}]^+$  produces  $[\text{Pd}(\text{Et}_4\text{dien})\text{OH}]^+$  and the rate increases at high pH. If the middle nitrogen has a  $\text{CH}_3$  group attached instead of a hydrogen atom, the rate of substitution is independent of pH. Explain these observations and the mechanism in the cases where H or  $\text{CH}_3$  groups are bonded to the middle nitrogen.
28. The vanadium complexes  $[\text{V}(\text{H}_2\text{O})_6]^{2+}$  and  $[\text{V}(\text{H}_2\text{O})_6]^{3+}$  undergo exchange of  $\text{H}_2\text{O}$  with solvent water,



but the reactions take place at very different rates. How would you explain this observation?

29. The hydrolysis of a series of complexes  $[\text{Co}(\text{NH}_3)_5\text{X}]^{2+}$ , where X is a halide ion, occurs with the rates varying with X as  $\text{I}^- > \text{Br}^- > \text{Cl}^- > \text{F}^-$ . Provide an explanation of this observed trend in rates of reactions.

# Organometallic Compounds

The chemistry of organometallic compounds has become one of the most important areas of chemistry, and it has erased any boundary that might have existed between inorganic and organic chemistry. A few types of organometallic compounds were described briefly in Chapters 8 and 10. However, because of their great utility, a more thorough treatment of the chemistry of these materials is warranted, and in this chapter, a more systematic overview of this area of chemistry will be presented.

If organometallic compounds are considered to be those in which a metal is bonded to an organic group, the range of materials is enormous. Such compounds include metal alkyls, metal–olefin complexes, Grignard reagents, metallocenes, and metal–carbonyl complexes. These important types of organometallic compounds and the reactions that they undergo will be the subject of this chapter.

It should also be mentioned at the outset that numerous organometallic compounds constitute hazardous materials. Several of them, especially those containing small alkyl groups bonded to reactive metals, are spontaneously flammable in air and react explosively with water. Others, especially those of mercury and cadmium, are extremely toxic. To make the situation worse, some of these compounds are readily absorbed through the skin. Needless to say, special precautions and laboratory techniques are necessary when working with any organometallic compounds that have these characteristics.

Although much of the chemistry of organometallics is of recent development, such compounds have been known for many years. Zeise's salt, a metal–olefin complex having the formula  $K[Pt(C_2H_4)Cl_3]$ , was prepared in 1825. The first metal alkyl has been known since 1849 when Frankland prepared diethyl zinc by the reaction



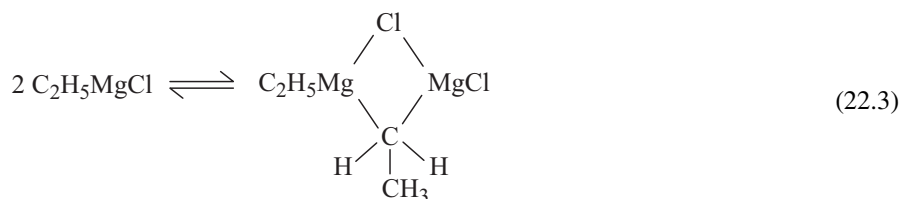
In 1900, Victor Grignard obtained the magnesium compounds now known as Grignard reagents by the general reaction

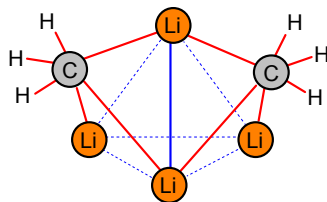


The growth of the general area of organometallic chemistry in the last century has been phenomenal. In a small book such as this, it is impossible to do more than give a general introduction to this important field, but more comprehensive surveys are listed in the references at the end of the chapter.

## 22.1 STRUCTURE AND BONDING IN METAL ALKYLs

Many metal alkyls display the characteristic of undergoing molecular association to give a variety of polymeric species. For example, Grignard reagents,  $\text{RMgX}$ , undergo association to give  $(\text{RMgX})_n$  in equilibrium processes that are both concentration and solvent dependent. In ether,  $\text{C}_6\text{H}_5\text{MgBr}$  gives an aggregate where  $n = 3$  in solutions as dilute as 1.0 molal. On the other hand,  $\text{C}_2\text{H}_5\text{MgBr}$  associates only slightly in tetrahydrofuran (THF) even at concentrations of 2–3 molal. Both the strength of the bonding in the aggregates and the solvating ability of the solvent are factors affecting the molecular association. When the solutions containing  $\text{RMgX}$  are evaporated, numerous  $\text{RMgX} \cdot x(\text{solvent})$  species are formed with an example being  $\text{C}_2\text{H}_5\text{MgBr} \cdot 2(\text{C}_2\text{H}_5)_2\text{O}$ . It is believed that association of  $\text{C}_2\text{H}_5\text{MgCl}$  involves the formation of a bridged species that can be shown as





**FIGURE 22.1** The structure of the methyl lithium tetramer. Only two methyl groups have been shown to simplify the figure and the blue lines outline the tetrahedral structure formed by the four lithium atoms.

Metal alkyls also form aggregates, and the association of  $\text{LiC}_2\text{H}_5$  has been investigated by the technique of mass spectrometry. The ions found in the mass spectrum consist of  $\text{Li}_n\text{R}_{n-1}^+$  with  $n = 1-6$ . However, the appearance potentials (which are associated with the ease of producing these ions) are considerably lower than for ions having other composition. Therefore, it was concluded that the equilibrium in the vapor is predominantly



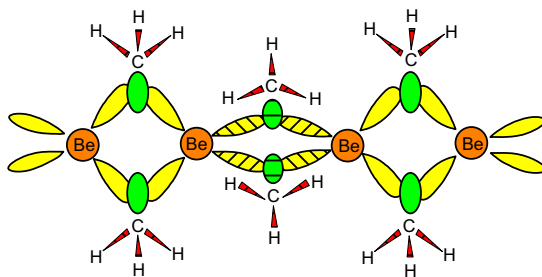
and that the other species are produced from fragmentation reactions. It has also been shown that the ratio of hexamer to tetramer decreases as the temperature increases. Molecular weight studies on benzene solutions of ethyllithium indicate that association occurs with the formation of  $(\text{LiC}_2\text{H}_5)_n$  with  $n$  in the range of about 4.5–6.0. The structure of the tetramer has the four lithium atoms arranged in a tetrahedron, whereas the structure of the hexamer has the six lithium atoms in a chair configuration of a six-membered ring. However, methyl lithium exists as the  $(\text{LiCH}_3)_4$  tetramer in both the solid state and in solutions with inert solvents. The structure of the tetramer is shown in Figure 22.1. The four lithium atoms reside at the corners of the tetrahedron, and a methyl group is located above the center of each triangular face. In the solid phase, the structure is body-centered cubic with a  $(\text{LiCH}_3)_4$  unit at each lattice site. Solid  $\text{LiC}_4\text{H}_9$  also consists of tetramers, and these units persist in solutions.

As has been shown earlier (see Chapter 9),  $\text{BH}_3$  dimerizes completely to  $\text{B}_2\text{H}_6$ . It is interesting that  $\text{B}(\text{CH}_3)_3$  does not dimerize and, in fact, the  $\text{BR}_3$  compounds in general are monomeric. On the other hand,  $\text{Be}(\text{CH}_3)_2$  is a polymeric material that has the structure shown in Figure 22.2. The  $\text{CH}_3$  bridges form three center bonds analogous to the  $\text{B}-\text{H}-\text{B}$  bonds in diborane. As was described in Chapter 8,  $\text{Mg}(\text{CH}_3)_2$  is similarly polymeric.

The extensive dimerization of  $\text{AlR}_3$  compounds is well known, and the  $[\text{Al}(\text{CH}_3)_3]_2$  dimer has the structure shown in Figure 22.3. Measurements of the molecular weights of aluminum alkyls in benzene solutions indicate that the methyl, ethyl, and  $n$ -propyl compounds exist as dimers. However, the heat of dissociation of the  $(\text{AlR}_3)_2$  dimers varies with the alkyl group as shown by the following values:

R =	CH <sub>3</sub>	C <sub>2</sub> H <sub>5</sub>	<i>n</i> -C <sub>3</sub> H <sub>7</sub>	<i>n</i> -C <sub>4</sub> H <sub>9</sub>	<i>i</i> -C <sub>4</sub> H <sub>9</sub>
$\Delta H_{\text{diss}}$ , kJ mol <sup>-1</sup>	81.2	70.7	87.4	38	33

As was described in Chapter 4, the solubility parameter,  $\delta$ , can often be used as a diagnostic tool for studying molecular association. Table 22.1 shows some of the relevant data for several aluminum alkyls. The solubility parameters were calculated from vapor pressure data using the procedure described in Chapter 4.



**FIGURE 22.2** The structure of polymeric  $(\text{Be}(\text{CH}_3)_2)_n$ . The two methyl groups in the middle of the drawing have the three hydrogen atoms directed toward the front and back of the page.



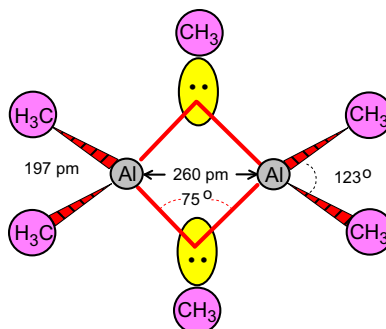


FIGURE 22.3 The structure of the dimer of trimethylaluminum.

TABLE 22.1 Physical Data for Some Aluminum Alkyls

Compound <sup>a</sup>	b.p., °C	$\Delta H_{\text{vap}}$ , kJ mol <sup>-1</sup>	$\Delta S_{\text{vap}}$ , J mol <sup>-1</sup> K <sup>-1</sup>	$\delta$ , J <sup>1/2</sup> cm <sup>-3/2</sup>
Al(CH <sub>3</sub> ) <sub>3</sub>	126.0	44.92	112.6	20.82
Al(C <sub>2</sub> H <sub>5</sub> ) <sub>3</sub>	186.6	81.19	176.6	23.75
Al( <i>n</i> -C <sub>3</sub> H <sub>7</sub> ) <sub>3</sub>	192.8	58.86	126.0	17.02
Al( <i>i</i> -C <sub>4</sub> H <sub>9</sub> ) <sub>3</sub>	214.1	65.88	135.2	15.67
Al(C <sub>2</sub> H <sub>5</sub> ) <sub>2</sub> Cl	208.0	53.71	111.6	19.88
Al(C <sub>2</sub> H <sub>5</sub> )Cl <sub>2</sub>	193.8	51.99	111.1	21.58

<sup>a</sup>Formula given is for the monomer.

In addition to the solubility parameter, the entropy of vaporization is a valuable piece of evidence in studying the association that occurs in liquids and vapors. In the case of [Al(C<sub>2</sub>H<sub>5</sub>)<sub>3</sub>]<sub>2</sub>, the entropy of vaporization (176.6 J mol<sup>-1</sup> K<sup>-1</sup>) is almost exactly *twice* the value of 88 J mol<sup>-1</sup> K<sup>-1</sup> predicted by Trouton's rule (which applies for random, unassociated liquids),

$$\Delta S_{\text{vap}} = \frac{\Delta H_{\text{vap}}}{T} \approx 88 \text{ J mol}^{-1} \text{ K}^{-1} \quad (22.5)$$

The value of 176.6 J mol<sup>-1</sup> K<sup>-1</sup> indicates that *one* mole of liquid is transformed into *two* moles of vapor when vaporization occurs. Therefore, it can be concluded that [Al(C<sub>2</sub>H<sub>5</sub>)<sub>3</sub>]<sub>2</sub> dimers are present in the liquid state, but the vapor consists of Al(C<sub>2</sub>H<sub>5</sub>)<sub>3</sub> monomers. Examination of the data for [Al(CH<sub>3</sub>)<sub>3</sub>]<sub>2</sub> shows an entropy of vaporization of 112.6 J mol<sup>-1</sup> K<sup>-1</sup>, which is substantially greater than the value of 88 J mol<sup>-1</sup> K<sup>-1</sup> predicted by Trouton's rule. This value could be interpreted as indicating that the liquid is only partially dimerized and that it is converted completely into monomers during vaporization. However, as will now be shown, this interpretation is incorrect.

The entropy of vaporization for [Al(CH<sub>3</sub>)<sub>3</sub>]<sub>2</sub> differs substantially from the value predicted by Trouton's rule, but from other evidence, it is known that the liquid contains dimers. However, if liquid trimethylaluminum contains dimers and the entropy of vaporization is lower than that expected if each dimer is converted into two molecules during vaporization, there must be some other explanation. Therefore, the entropy of vaporization being higher than expected for a liquid that behaves normally may indicate a liquid that is completely dimerized that is only partially dissociated during vaporization.

Because triethylaluminum fits clearly the case of a dimerized liquid that dissociates completely into monomers in the vapor, other indicators are needed to apply in the trimethylaluminum case. One type of evidence is provided by the solubility parameters that have values of 20.8 and 23.7 J<sup>1/2</sup> cm<sup>-3/2</sup> for trimethylaluminum and triethylaluminum, respectively. These values are quite similar and in the expected order for compounds differing slightly in molecular masses. Because the solubility parameters do not differ significantly, it appears that both liquids must exist in the same form in the liquid state. Therefore, because the entropy of vaporization of trimethylaluminum is not twice the value predicted by Trouton's rule (as is that for triethylaluminum), it can be concluded that trimethylaluminum is a completely

dimerized liquid that only partially dissociates to monomers during vaporization. The reason for this difference is that the boiling point of triethylaluminum is 186.6 °C whereas trimethylaluminum boils at 126.0 °C. This difference is sufficient to cause complete dissociation of  $[\text{Al}(\text{C}_2\text{H}_5)_3]_2$  during vaporization but only partial dissociation of  $[\text{Al}(\text{CH}_3)_3]_2$  at its lower boiling point.

Support for this conclusion is found by considering the cases of  $\text{Al}(n\text{-C}_3\text{H}_7)_3$  and  $\text{Al}(i\text{-C}_4\text{H}_9)_3$ . The entropies of vaporization for these compounds are 126.0 and 135.2  $\text{J mol}^{-1} \text{K}^{-1}$ , respectively. These values are larger than the value of 88  $\text{J mol}^{-1} \text{K}^{-1}$  predicted by Trouton's rule, so a change in molecular association during vaporization is involved. The solubility parameters of 17.0 and 15.7  $\text{J}^{1/2} \text{cm}^{-3/2}$  for  $\text{Al}(n\text{-C}_3\text{H}_7)_3$  and  $\text{Al}(i\text{-C}_4\text{H}_9)_3$  are considerably *smaller* than are the values for the methyl and ethyl compounds. Because the solubility parameter reflects the cohesion energy of the liquid, it can be concluded that  $\text{Al}(n\text{-C}_3\text{H}_7)_3$  and  $\text{Al}(i\text{-C}_4\text{H}_9)_3$  are only *partially* dimerized in the liquid phase. Thus, the entropies of vaporization being higher than predicted from Trouton's rule is the result of partially dimerized liquids being converted into vapors that are completely monomeric. Both compounds have boiling points (192.8 and 214.1 °C for  $\text{Al}(n\text{-C}_3\text{H}_7)_3$  and  $\text{Al}(i\text{-C}_4\text{H}_9)_3$ , respectively) that are even higher than that of  $[\text{Al}(\text{C}_2\text{H}_5)_3]_2$ , which completely dissociates during vaporization. Therefore, both  $[\text{Al}(n\text{-C}_3\text{H}_7)_3]_2$  and  $[\text{Al}(i\text{-C}_4\text{H}_9)_3]_2$  should be completely dissociated in the vapor phase also and only partially dimerized in the liquid state. The data shown earlier for the heat of dissociation of the dimers also support this interpretation.

From the discussion above, the behavior of the lower aluminum alkyls can be summarized as shown below.

Monomer Formula				
	$\text{Al}(\text{CH}_3)_3$	$\text{Al}(\text{C}_2\text{H}_5)_3$	$\text{Al}(n\text{-C}_3\text{H}_7)_3$	$\text{Al}(i\text{-C}_4\text{H}_9)_3$
Liquid	Dimers	Dimers	Dimers + monomers	Dimers + monomers
Vapor	Dimers + monomers	Monomers	Monomers	Monomers

These observations are for the pure compounds, but the association of alkyl aluminum compounds also occurs in solution. As described earlier in connection with the association of Grignard reagents, the association metal alkyls often gives different types of aggregates depending on the nature of the solvent.

Other types of association are found with other metal alkyls. For example,  $\text{Ga}(\text{CH}_3)_3$  is a dimer that dissociates upon vaporization, but the analogous compounds containing larger alkyl groups exist as monomers. The low solubility parameters of 13.6  $\text{J}^{1/2} \text{cm}^{-3/2}$  and 18.2  $\text{J}^{1/2} \text{cm}^{-3/2}$  for  $\text{Ge}(\text{CH}_3)_4$  and  $\text{Ge}(\text{C}_2\text{H}_5)_4$ , respectively, indicate that these compounds are not associated in the liquid phase. Molecular association of metal alkyls occurs more often when the compound can be considered as being electron deficient (e.g.,  $\text{AlR}_3$ ,  $\text{BeR}_2$ , etc.).

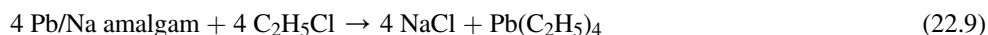
## 22.2 PREPARATION OF ORGANOMETALLIC COMPOUNDS

Because of the wide range of organometallic compounds, a large number of reactions have been employed in their synthesis. A few of the preparative methods for compounds of the Group IA and IIA metals were shown in Chapter 8. In this section, some of the types of synthetic reactions will be shown with an emphasis on methods that have general applicability.

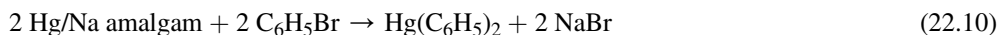
1. *The reaction of a metal with an alkyl halide.* This technique is widely used when the metal is an active one. The equations will be written as if monomeric species are formed although, as has been discussed, some of the metal alkyls are associated. Examples of this type of reaction are the following.



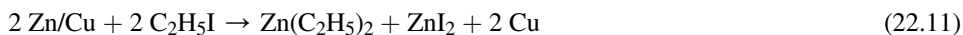
Before its use in motor fuels was banned in the United States, tetraethyl lead was made in large quantities. It was produced by a reaction in which lead was made more reactive by its amalgamation with sodium, and the amalgam was reacted with ethyl chloride.



A similar reaction can be carried out in which mercury is amalgamated with sodium.



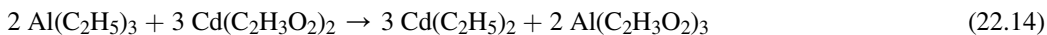
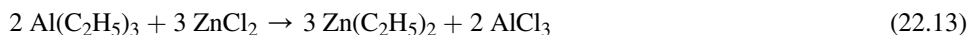
The driving force for these reactions is the formation of an ionic sodium salt. A variation of this approach makes use of the fact that one metal causes the more rapid oxidation of another when the two are in contact because of the electrical potential of the couple. A reaction of this type is the following.



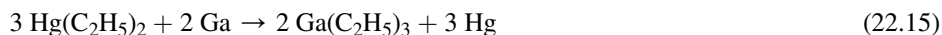
2. *Alkyl group transfer.* When metal alkyls react with compounds of a different metal, the products frequently are those in which one product forms a stable crystal lattice. For example, in the reaction



NaCl forms a stable lattice. This reaction can be considered by the hard–soft interaction principle (see Chapter 6) to occur as a result of the favorable interaction of  $\text{Na}^+$  with  $\text{Cl}^-$ . Therefore, the lattice energy of one product becomes the driving force for the metathesis reaction. Other examples of this type of reaction are the following.



The dialkylmercury compounds are useful reagents for preparing a large number of alkyls of other metals by group transfer reactions. This is especially true because mercury(II) is easily reduced. The following reactions are examples of this type of reaction.



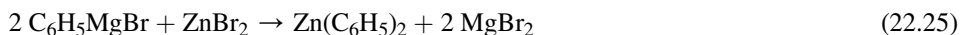
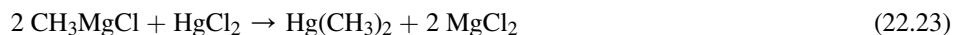
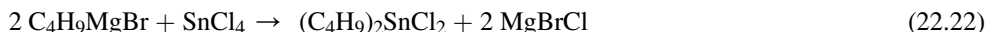
In the last of these reactions, low boiling hydrocarbons are used as solvents and because the sodium alkyls are predominantly ionic, they are insoluble in such solvents. Alkyls of other Group IA metals (shown as M in the following equation) can also be produced by this type of reaction, and benzene is a frequently used solvent.



3. *Reaction of a Grignard reagent with a metal halide.* Grignard reagents are prepared by the general reaction

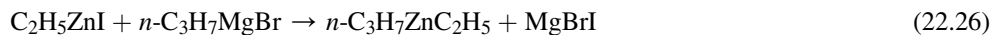


These compounds function as alkyl group transfer agents, particularly when the other compound is a somewhat covalent halide. Some typical reactions that illustrate alkyl group transfer are the following



Note that in these equations,  $2 \text{ MgBrCl}$  is equivalent to  $\text{MgBr}_2 + \text{MgCl}_2$  regardless of whether the product is a mixed halide or a mixture of the two halides. Thus, formulas such as  $\text{MgBrCl}$  will continue to be used to simplify the equations.

The reaction



yields a product of the type  $\text{RZnR}'$ , but in this case the product is converted on standing to  $\text{ZnR}_2$  and  $\text{ZnR}'_2$ .

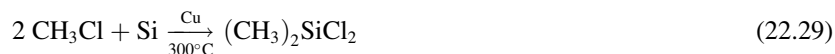


Compounds such as  $\text{R}_2\text{SnCl}_2$  and  $\text{R}_2\text{SiCl}_2$  have several uses because two reactive Sn—Cl or Si—Cl bonds remain that can undergo hydrolysis, reactions with alcohols, etc. The reactions of Grignard reagents constitute one of the most general synthetic procedures, and they can be used in a wide variety of cases.

4. *Reaction of an olefin with hydrogen and a metal.* In some cases, it is possible to synthesize a metal alkyl directly from the metal. A reaction of this type can be used to prepare aluminum alkyls.



A different type of direct synthesis involves the reaction of the metal with an alkyl halide.



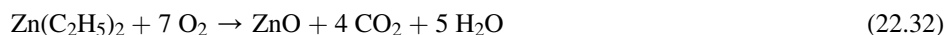
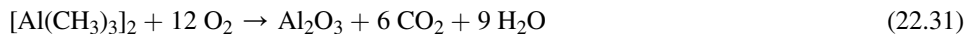
5. *Substitution of alkyl groups.* In some cases, an organic group will replace a different alkyl group from an organo-metallic compound. For example, the reaction of ethyl sodium and benzene can be shown as follows.



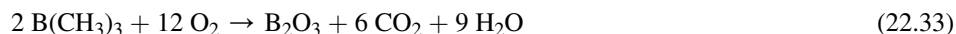
Although many other synthetic routes are utilized in specific cases, the methods described have been used extensively.

## 22.3 REACTIONS OF METAL ALKYLs

Most metals are extremely reactive when they are in the form of powders, and they may react explosively with oxygen, halogens, sulfur, or other oxidizing agents. In some ways, metal alkyls behave as if they contained the metals in an atomic form. As a result, some metal alkyls are spontaneously flammable in air as shown in the following cases.

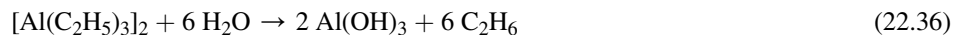


Those that are not spontaneously flammable will burn readily as a result of the formation of very stable combustion products. For example, the combustion of trimethylboron produces three stable oxides.

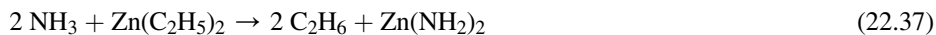


Virtually all reactions of this type are extremely exothermic. However, tetraethyl lead,  $\text{Pb}(\text{C}_2\text{H}_5)_4$ , is stable in air as are  $\text{SiR}_4$ ,  $\text{GeR}_4$ , and  $\text{HgR}_2$  compounds, although they will all readily burn.

Most metal alkyls react vigorously (some explosively) with water to produce a hydrocarbon and a metal hydroxide. Typical reactions are the following.



In these reactions, the polar nature of the metal—carbon bond is indicated by the fact that the electronegativity of carbon is approximately 2.5 whereas that of most metals is in the range 1.0–1.5. Thus, many reactions take place as if  $\text{R}^-$  is present, with the alkyl group behaving as a strong base. This behavior extends even to reactions with  $\text{NH}_3$ , which is normally basic, but an amide is formed as a result the removal of  $\text{H}^+$ .



The lithium alkyls tend to behave as if they were more covalent than those of the other alkali metals. Two of the reasons for this are the higher ionization potential of lithium and the more favorable overlap of its smaller 2s orbital with a carbon orbital leading to a bond that is more covalent. As a result, the lithium alkyls are generally more soluble in hydrocarbon solvents than are those of the other metals in Group IA. It is interesting to note that the greater ionic character of the other alkali metal compounds is also indicated by the fact that at room temperature, butyllithium is a liquid but butylsodium is a solid. As a general rule, the sodium alkyls are more reactive because they behave as if they contain a negative carbanion.

Although the alkyl compounds of B, Hg, and Tl do not react readily with water, they do undergo reactions with acids.



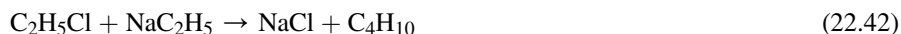
In the latter case, a series of reactions is involved with only the first step being shown. Although the methyl groups in  $\text{Ga}(\text{CH}_3)_3$  are not held by ionic bonds, they still behave as proton acceptors as shown by the reaction



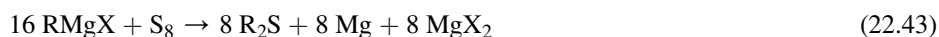
Metal alkyls also undergo a variety of reactions with nonmetallic elements, and the reactions with oxygen have already been described. For example, reactions with halogens produce an alkyl halide and the metal halide.



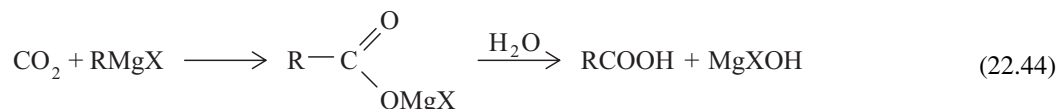
Metal alkyls will also react with alkyl halides with the formation of a hydrocarbon and a metal salt.



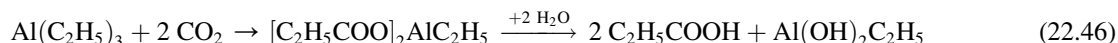
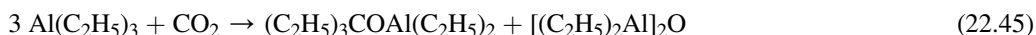
Grignard reagents undergo many reactions, and the reaction with sulfur can be shown as follows.



Some metal alkyls will react with carbon dioxide, and the reaction of Grignard reagents with  $\text{CO}_2$  can be used to prepare an acid.



Aluminum alkyls (written here as if they were monomeric) can react with  $\text{CO}_2$  in two ways, which are represented by the following equations.



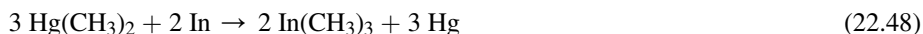
However, not all metal alkyls will react with  $\text{CO}_2$ . As a general rule, a metal alkyl will react with  $\text{CO}_2$  if the metal has an electronegativity below about 1.5. Although it has been mentioned before, when burning magnesium is thrust into a bottle of  $\text{CO}_2$  it continues to burn.



The reactions shown above indicate that  $\text{CO}_2$  is also not an inert atmosphere toward many metal alkyls.

Although a brief survey of several types of reactions of metal alkyls has been given, it is the transfer of organic groups that is most important. Of course, some of the reactions already shown are of this type, but there is a very extensive use of these compounds in organic synthesis. Only a very brief survey of some of the reactions will be shown here, and the suggested readings at the end of this chapter should be consulted for more complete coverage of this important topic.

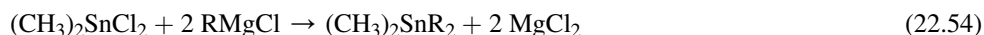
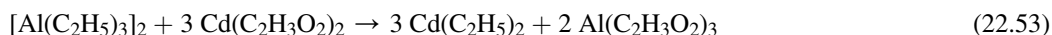
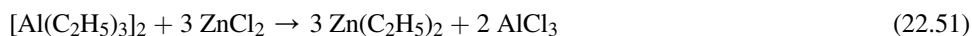
One useful type of reaction of metal alkyls is the transfer of an alkyl group to another metal as shown in the following examples.



The transfer of the alkyl group may be to another metal that is already contained in a compound as shown in the following equation.



The product in this case contains a three-membered ring of Sn atoms (see Chapter 12) each of which is bonded to two R groups. Other reactions of this type are shown below.

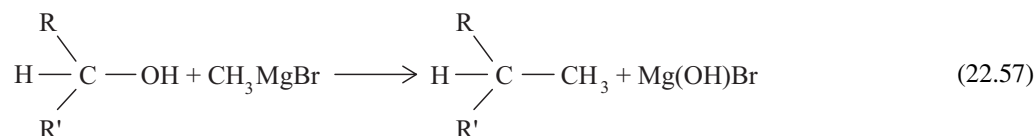


Reactions such as those shown above are useful in the preparation of an enormous range of organometallic compounds.

Perhaps the greatest use of organometallic compounds is in the transfer of alkyl groups to carbon atoms that are bound in organic molecules. For example, the reaction of a Grignard reagent with a primary alcohol can be shown as



With a secondary alcohol, the reaction is



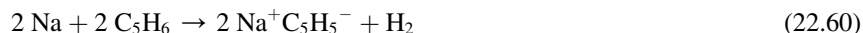
Metal alkyls also react with alcohols as illustrated in the following reactions.



The reactions shown above illustrate the transfer of alkyl groups, and a large number of reactions of this type are known.

## 22.4 CYCLOPENTADIENYL COMPLEXES (METALLOCENES)

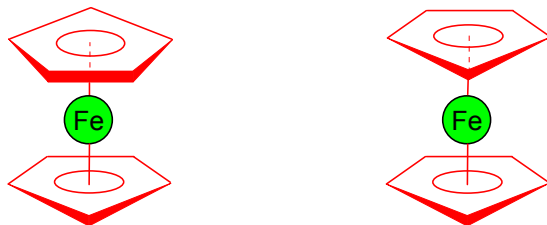
When an active metal such as sodium or magnesium reacts with cyclopentadiene ( $\text{C}_5\text{H}_6$ ), a compound is formed that contains the cyclopentadienyl anion,  $\text{C}_5\text{H}_5^-$ .



When a compound of this type reacts with  $\text{FeCl}_2$ , the reaction is



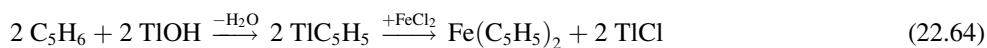
In  $\text{Fe}(\text{C}_5\text{H}_5)_2$  (also written sometimes as  $\text{Fe}(\text{cp})_2$  and known as dicyclopentadienyliron or *ferrocene*), the  $\text{C}_5\text{H}_5^-$  rings behave as six electron donors. Because  $\text{Fe}^{2+}$  has 24 electrons, gaining 12 more from two  $\text{C}_5\text{H}_5^-$  ligands gives 36 electrons around the Fe, which gives it the configuration of the next noble gas. In other words,  $\text{Fe}(\text{C}_5\text{H}_5)_2$  obeys the 18-electron rule (see Section 22.6). The structure of ferrocene can be shown as either a staggered (left) or eclipsed (right) configuration. For obvious reasons, this structure is sometimes referred to as a “sandwich” compound. The barrier to ring rotation is approximately 4.4 kJ/mol so rotation due to thermal energy is possible at normal temperatures. However, solid ferrocene undergoes a phase transition at 164 K and the barrier to rotation appears to be different for the two phases probably as a result of *intermolecular* forces in the solid crystal.



Following the original discovery of ferrocene in 1951, cyclopentadienyl complexes of other first row transition metals were prepared. These include  $\text{Ti}(\text{C}_5\text{H}_5)_2$ ,  $\text{Cr}(\text{C}_5\text{H}_5)_2$ ,  $\text{Mn}(\text{C}_5\text{H}_5)_2$ ,  $\text{Co}(\text{C}_5\text{H}_5)_2$ , and  $\text{Ni}(\text{C}_5\text{H}_5)_2$ . Several methods of preparation have been useful for synthesizing compounds of this type. In one of these methods, a Grignard reagent is prepared containing the  $\text{C}_5\text{H}_5$  group, which then reacts with a metal halide.

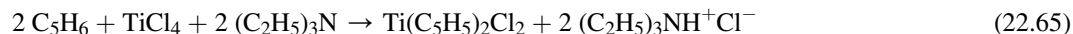


where M is a transition metal. Other methods for preparing metallocenes depend on the generation of the  $\text{C}_5\text{H}_5^-$  ion by a base in the presence of a metal to which it bonds. Reactions of this type include the following.



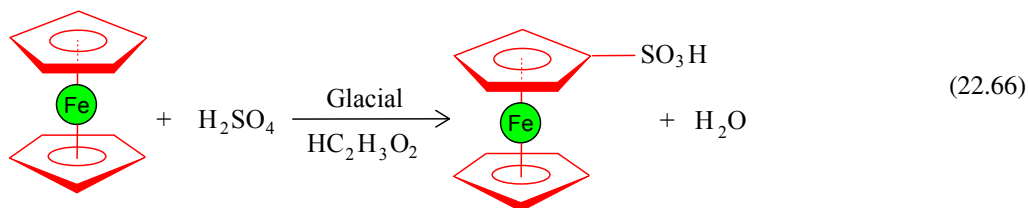
These reactions depend on the basicity of the nonaqueous solvent  $(\text{C}_2\text{H}_5)_2\text{NH}$  or the  $\text{TIOH}$  (a strong base), respectively.

Another interesting cyclopentadienyl complex is  $\text{Ti}(\text{C}_5\text{H}_5)_2\text{Cl}_2$ , which is prepared by a reaction that is analogous to that shown in Eq. (22.63). The reaction can be shown as follows.

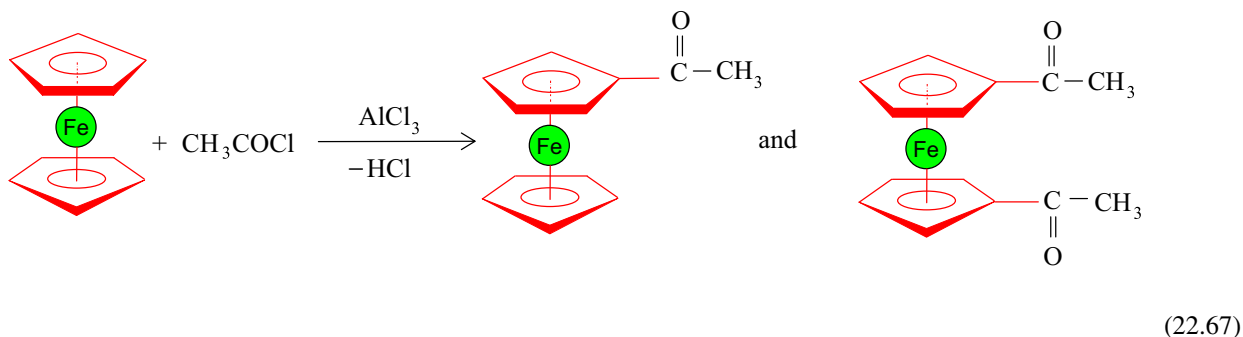


A very large number of compounds containing the  $\text{Ti}(\text{C}_5\text{H}_5)_2^{2+}$  group have been obtained by reactions of  $\text{Ti}(\text{C}_5\text{H}_5)_2\text{Cl}_2$ , which contains two reactive Ti–Cl bonds.

Because of the stability of ferrocene (it melts at  $173^\circ\text{C}$  and it is stable to approximately  $500^\circ\text{C}$ ), many organic reactions can be carried out on the rings without disruption of the complex. It can be sulfonated as shown in the reaction

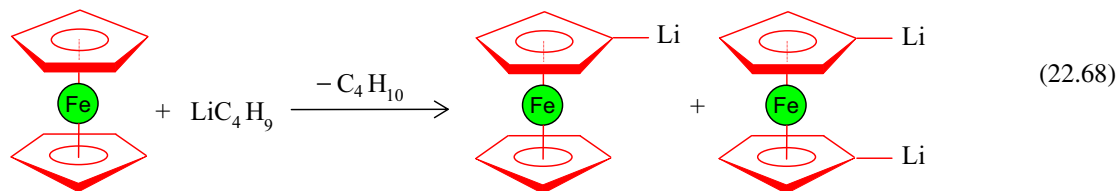


Ferrocene will undergo Friedel–Crafts reactions to give mono- and disubstituted acylated derivatives.



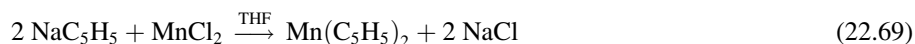
and it is even possible to connect the rings in ferrocene by a  $-\text{CH}_2\text{CH}_2\text{CH}_2-$  linkage.

The reaction of ferrocene with butyllithium produces reactive intermediates that can be used to prepare numerous other derivatives.

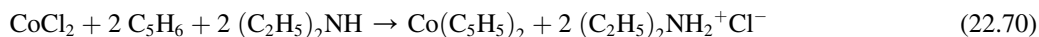


The ferrocene derivatives containing lithium are the starting compounds for attaching  $-\text{NO}_2$ ,  $-\text{COOH}$ ,  $-\text{B}(\text{OH})_2$ ,  $-\text{NH}_2$ , and other substituents to the  $\text{C}_5\text{H}_5$  rings.

In addition to the extensive chemistry of ferrocene, other metallocenes undergo some of the same types of reactions. Manganocene,  $\text{Mn}(\text{C}_5\text{H}_5)_2$ , is prepared by the following reaction that is carried out in THF as a solvent.



Cobaltocene can be prepared from  $\text{CoCl}_2$  and  $\text{C}_5\text{H}_6$  in a one-step procedure that can be carried out in diethylamine solution,



in which the loss of  $\text{H}^+$  from  $\text{C}_5\text{H}_6$  is assisted by the solvent, which functions as a base.

Literally thousands of derivatives of the metallocenes have been prepared, and their study constitutes a growing field of inorganic chemistry. Only a cursory survey can be presented here, and the references given at the end of this chapter should be consulted for treatment at a higher level.

## 22.5 METAL CARBONYL COMPLEXES

Metal carbonyl complexes are an interesting series of coordination compounds in which the ligands are CO molecules and in many cases, the metals are present in a zero oxidation state. In these complexes, both the metal and ligand are soft according to the Lewis acid–base definitions. Although the discussion at first will be limited to the binary compounds containing only metal and CO, many mixed complexes are known that contain both CO and other ligands.

### 22.5.1 Binary Metal Carbonyls

The first metal carbonyl prepared was  $\text{Ni}(\text{CO})_4$ , and it was obtained by L. Mond in the 1890s. This extremely toxic compound was prepared by first reducing nickel oxide with hydrogen,



and then treating the Ni with CO,



Cobalt does not react with CO under these conditions. Because  $\text{Ni}(\text{CO})_4$  is volatile (b.p.  $43^\circ\text{C}$ ), this procedure affords a method for separating Ni from Co by the process now known as the *Mond process*. Although many complexes are known that contain both carbonyl and other ligands (the so-called mixed carbonyl complexes), the number containing only a metal and carbonyl ligands (the binary metal carbonyls) is rather small, and they are listed in [Table 22.2](#).

The composition of most stable binary metal carbonyls can be predicted by the effective atomic number rule (EAN) or the 18-electron rule as it is also known. The idea underlying the rule is that a metal in the zero or other low oxidation state will gain electrons from a sufficient number of ligands to allow the metal to achieve the electron configuration of the next noble gas. For the first row transition metals, this means the krypton configuration with a total of 36 electrons, which gives the outer shells filled with 18 electrons. Thus, the transition metal acquires the electron configuration of krypton, so its so-called EAN is 36. Although the terminologies EAN and 18-electron rule are both in common use, the former will be used in this chapter.



**TABLE 22.2** Binary Metal Carbonyls

Mononuclear Compound	m.p., °C	Dinuclear Compound	m.p., °C	Polynuclear Compound	m.p. °C
Ni(CO) <sub>4</sub>	-25	Mn <sub>2</sub> (CO) <sub>10</sub>	155	Fe <sub>3</sub> (CO) <sub>12</sub>	140 (d)
Fe(CO) <sub>5</sub>	-20	Fe <sub>2</sub> (CO) <sub>9</sub>	100 (d)	Ru <sub>3</sub> (CO) <sub>12</sub>	—
Ru(CO) <sub>5</sub>	-22	Co <sub>2</sub> (CO) <sub>8</sub>	51	Os <sub>3</sub> (CO) <sub>12</sub>	224
Os(CO) <sub>5</sub>	-15	Rh <sub>2</sub> (CO) <sub>8</sub>	76	Co <sub>4</sub> (CO) <sub>12</sub>	60 (d)
Cr(CO) <sub>6</sub>	Subl.	Tc <sub>2</sub> (CO) <sub>10</sub>	160	Rh <sub>4</sub> (CO) <sub>12</sub>	150 (d)
Mo(CO) <sub>6</sub>	Subl.	Re <sub>2</sub> (CO) <sub>10</sub>	177	Ir <sub>4</sub> (CO) <sub>13</sub>	210 (d)
V(CO) <sub>6</sub>	70 (d)			Rh <sub>6</sub> (CO) <sub>16</sub>	200 (d)
W(CO) <sub>6</sub>	Subl.				

For complexes of metals in the zero oxidation state containing soft ligands such as CO, PR<sub>3</sub>, alkenes, etc., there is a strong tendency for the stable complexes to be those containing the number of ligands predicted by the EAN rule. Because the Ni atom has 28 electrons, it would be expected to gain 8 additional electrons from 4 ligands each donating a pair of electrons. Thus, the stable compound formed between nickel and carbon monoxide is Ni(CO)<sub>4</sub>. Iron contains 26 electrons, so bonding to 5 CO ligands brings the total to 36, and the stable binary carbonyl is Fe(CO)<sub>5</sub>. For chromium(0), which contains 24 electrons, the stable carbonyl is Cr(CO)<sub>6</sub> as expected on the basis of the EAN rule. The 6 pairs of electrons from the 6 CO ligands bring to 36, the number of electrons around the chromium atom.

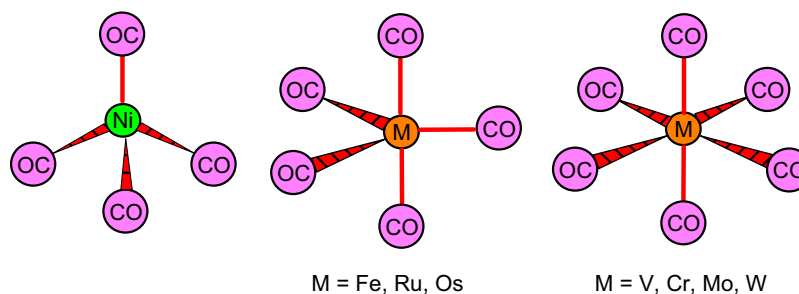
The Mn atom has 25 electrons. Adding 5 carbonyl groups raises the number of electrons to 35, leaving the Mn atom 1 electron short of the krypton configuration. If the unpaired electron on one manganese atom that has five CO molecules attached is then allowed to pair up with an unpaired electron on another Mn atom with five CO ligands to form a metal–metal bond, the resulting formula is (CO)<sub>5</sub>Mn–Mn(CO)<sub>5</sub> or [Mn(CO)<sub>5</sub>]<sub>2</sub>, which is the simplest formula for a binary manganese carbonyl obeying the EAN rule.

Cobalt has 27 electrons so adding 8 electrons from 4 CO ligands gives a total of 35 electrons on Co. Forming a metal–metal bond using the single unpaired electron from each cobalt atom bound to four CO molecules would give [Co(CO)<sub>4</sub>]<sub>2</sub> or Co<sub>2</sub>(CO)<sub>8</sub> as the stable carbonyl compound. In this case, there are two different structures possible for Co<sub>2</sub>(CO)<sub>8</sub> that obey the EAN rule. Based on the EAN rule, Co(CO)<sub>4</sub> should not be stable, but the species Co(CO)<sub>4</sub><sup>-</sup> is stable because a total of nine electrons have been added to Co.

Metal carbonyls are usually named by giving the name of the metal followed by the ligand name, carbonyl, preceded by a prefix to indicate the number of CO groups. For example, Ni(CO)<sub>4</sub> is nickel tetracarbonyl, Cr(CO)<sub>6</sub> is chromium hexacarbonyl, and Fe<sub>2</sub>(CO)<sub>9</sub> is diiron nonacarbonyl.

### 22.5.2 Structures of Metal Carbonyls

Structures of the mononuclear (containing only one metal atom) carbonyls are comparatively simple. As shown in Figure 22.4, nickel tetracarbonyl is tetrahedral, the pentacarbonyls of iron, ruthenium, and osmium are trigonal bipyramidal, and the hexacarbonyls of vanadium, chromium, molybdenum, and tungsten are octahedral.



**FIGURE 22.4** The structures of mononuclear carbonyls. The bonds are between the metal and the carbon atom in CO in all cases.

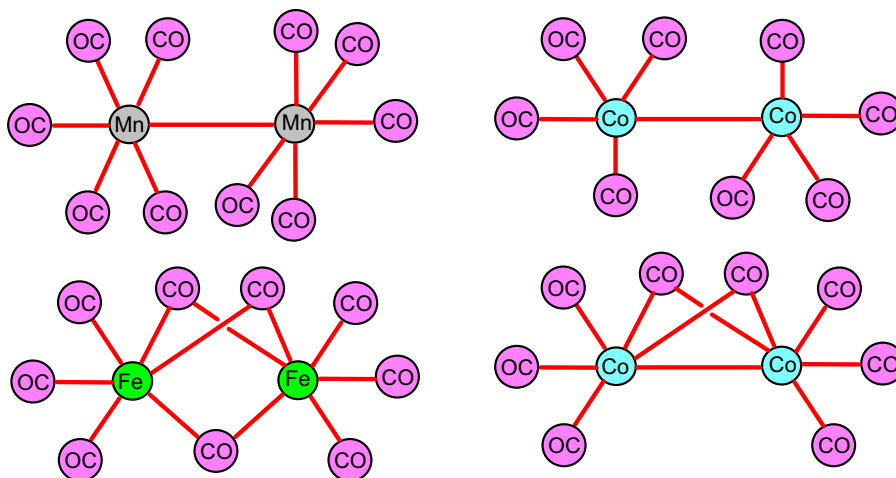


FIGURE 22.5 The structures of some dinuclear metal carbonyls including  $[\text{Co}(\text{CO})_4]_2$  that exists in two forms that undergo rapid equilibration.

Dinuclear metal carbonyls contain two metal atoms and involve either metal–metal bonds or bridging CO groups, or both. For example, the structure of  $\text{Fe}_2(\text{CO})_9$ , diiron nonacarbonyl, contains three CO ligands that form bridges between the iron atoms as well as three other CO groups bonded only to one atom. Carbonyl groups that are attached to two metal atoms simultaneously are called *bridging* carbonyls whereas those attached to only one metal atom are referred to as *terminal* carbonyl groups. The structures of  $\text{Mn}_2(\text{CO})_{10}$ ,  $\text{Tc}_2(\text{CO})_{10}$ , and  $\text{Re}_2(\text{CO})_{10}$  actually involve only a metal–metal bond so the formulas are more correctly written as  $(\text{CO})_5\text{M}-\text{M}(\text{CO})_5$ . Two isomers are known for  $\text{Co}_2(\text{CO})_8$ . One has a metal–metal bond between the cobalt atoms whereas the other has two bridging CO ligands and a metal–metal bond. Figure 22.5 shows the structures of the dinuclear metal carbonyls. In  $\text{Mn}_2(\text{CO})_{10}$ , the Mn atoms do not lie in the same plane as the four CO ligands. The CO ligands lie about 12 pm out of the plane containing the Mn on the side toward the fifth CO group. The tri- and tetranuclear carbonyls can be considered as clusters of metal atoms containing metal–metal bonds or bridging carbonyls or both as shown in Figure 22.6.

The nature of the bonds between metals and carbon monoxide is interesting and deserves some discussion. Carbon monoxide has the valence bond structure shown as



with a triple bond between C and O. The formal charge on the oxygen atom is +1 whereas that on the carbon atom is –1. The carbon end of the CO molecule is thus a better electron pair donor, and it is the carbon atom that bonds to the metal.

Figure 22.7 shows the molecular orbital energy level diagram for CO, which shows that the bond order for the molecule is 3. For gaseous CO, the C–O stretching band is observed at  $2143 \text{ cm}^{-1}$ . Typically, in metal carbonyls the C–O stretching band is seen at  $2000\text{--}2050 \text{ cm}^{-1}$  for terminal CO groups. The shift of the CO stretching band upon coordination to metals reflects a slight reduction in the bond order resulting from back donation of electron density from

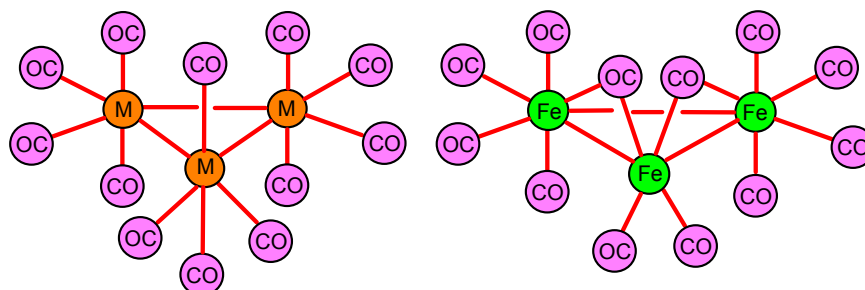


FIGURE 22.6 The structures  $\text{Ru}_3(\text{CO})_{12}$  and  $\text{Os}_3(\text{CO})_{12}$  (left) and  $\text{Fe}_3(\text{CO})_{12}$  (right).

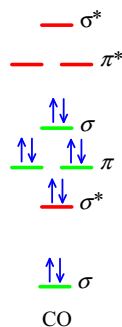
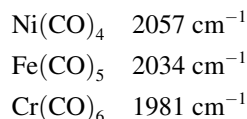


FIGURE 22.7 The molecular orbital diagram for CO.

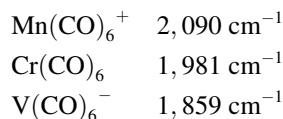
the metal to the antibonding orbitals on CO. This occurs to relieve part of the negative formal charge that occurs when the metal atom accepts pairs of electrons from the CO ligands. This occurs because the  $\pi^*$  orbital on CO has the appropriate symmetry to interact with a nonbonding  $d$  orbital on the metal. This interaction can be illustrated as shown in Figure 22.8.

Back donation occurs to an extent that varies depending on the nature of the metal and the number of ligands. For example, in  $\text{Fe}(\text{CO})_5$ , the formal charge on iron is  $-5$ , and in  $\text{Cr}(\text{CO})_6$  the formal charge on chromium is  $-6$ . Therefore, back donation is more extensive in both of these compounds than it is in the case of  $\text{Ni}(\text{CO})_4$  in which the formal charge on Ni is  $-4$ . Because the greater extent of back donation results in a greater reduction in C–O bond order, the infrared spectra of these compounds should show this effect as a shift in the position of the band corresponding to stretching the CO bond. The positions of the bands in the infrared spectra corresponding to the CO stretching vibration in these compounds are as follows.



Clearly, the CO bond is weakened to a greater extent as the degree of back donation increases.

A similar reduction in C–O bond order occurs for an isoelectronic series of metal carbonyls as the oxidation states of the metals are reduced. Consider the positions of the C–O stretching bands in the following carbonyls.



In this series, the metal has a progressively greater negative charge, which is partially relieved by back donation, and that results in a shift of the CO stretching vibration to lower wave numbers. However, there is a corresponding increase in metal–carbon stretching frequency showing the increased tendency to form multiple metal–carbon bonds to relieve part of the negative charge.

Stretching bands for bridging CO ligands are found in the region of the infrared spectrum characteristic of ketones, about  $1700\text{--}1800 \text{ cm}^{-1}$ . Accordingly, the infrared spectrum of  $\text{Fe}_2(\text{CO})_9$  shows absorption bands at  $2000 \text{ cm}^{-1}$  (terminal carbonyl stretching) and  $1830 \text{ cm}^{-1}$  (bridging carbonyl stretching). In most cases of bridging carbonyls, the C–O stretching band is seen around  $1850 \text{ cm}^{-1}$ . Although the effect of back donation on the CO bonds is easily explained, it

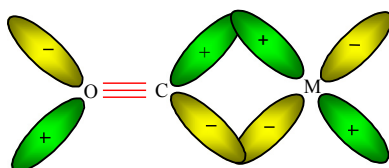


FIGURE 22.8 The overlap of a  $d$  orbital on a metal with the  $\pi^*$  orbital on CO leading to back donation.

should be pointed out that the spectra of metal carbonyls are frequently complex as a result of absorptions arising from other vibrational changes and combinations of bands.

### 22.5.3 Preparation of Metal Carbonyls

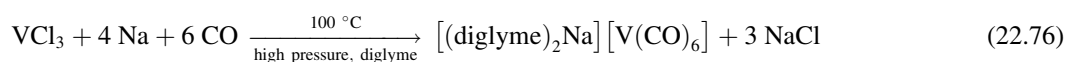
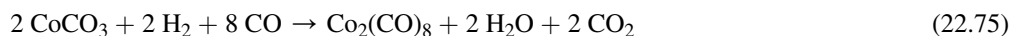
As has been described previously, Ni(CO)<sub>4</sub> can be prepared directly by the reaction of nickel with carbon monoxide, but most of the binary metal carbonyls listed in Table 22.2 cannot be obtained by this type of reaction. Other preparative techniques have been used to prepare metal carbonyls, and a few general procedures will be described here.

1. *Reaction of a metal with carbon monoxide.* The reactions with Ni and Fe proceed rapidly at low temperature and pressure.



For some other metals, high temperatures and pressures are required for the metal to react with CO. By using direct combination reactions, Co<sub>2</sub>(CO)<sub>8</sub>, Mo(CO)<sub>6</sub>, Ru(CO)<sub>5</sub>, and W(CO)<sub>6</sub> have been prepared when suitable conditions are used.

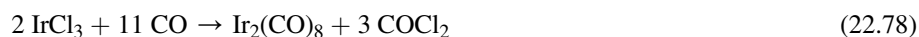
2. *Reductive carbonylation.* In this type of reaction, a metal compound is reduced in the presence of CO. Several types of reducing agents may be used depending on the particular synthesis being carried out. As illustrated by the following reactions, typical reducing agents are hydrogen, LiAlH<sub>4</sub>, and metals.



In the latter case, V(CO)<sub>6</sub> is obtained by hydrolyzing the product with H<sub>3</sub>PO<sub>4</sub> and subliming V(CO)<sub>6</sub> at 45–50 °C. The reduction of CoI<sub>2</sub> can be used to prepare Co<sub>2</sub>(CO)<sub>8</sub>.



3. *Displacement reactions.* The carbonyls of some metals have been prepared by the reaction of metal compounds directly with CO because CO is a reducing agent as well as a ligand. The following equations illustrate this type of process.

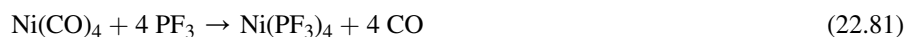
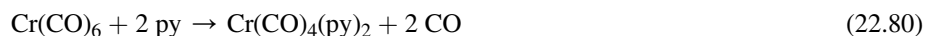


In such reactions, it *appears* as if CO has displaced a halogen or oxygen, but of course the reactions are more complex, and they involve the reduction of the metal.

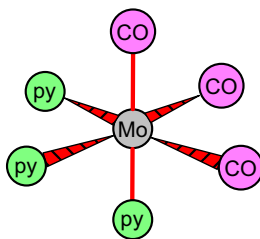
### 22.5.4 Reactions of Metal Carbonyls

Because metal carbonyls are reactive compounds, it is possible to carry out a large number of reactions that lead to useful derivatives. A few of the more general and important types of reactions will be described in this section.

1. *Substitution reactions.* Metal carbonyls undergo many reactions to produce mixed carbonyl complexes by the replacement of one or more carbonyl groups in a substitution reaction. The following equations show examples of substitution reactions of carbonyls.



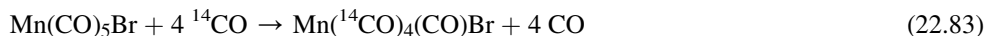
The structure of  $\text{Mo}(\text{CO})_3(\text{py})_3$  has all three CO ligands *trans* to pyridine ligands so the resulting structure is



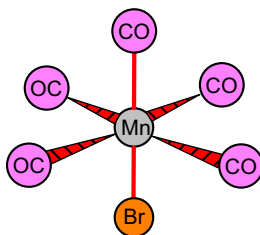
In Chapter 21, the *trans* effect exerted by ligands in square planar complexes was shown to be related to the ability of the ligands to form  $\pi$  bonds to the metal. Although the existence of a *trans* effect in octahedral complexes is much less clear than for square planar complexes, it nonetheless exists. It is known that CO is a ligand that gives rise to a substantial *trans* effect. Therefore, each CO in  $\text{Mo}(\text{CO})_6$  exerts an effect on the CO *trans* to it. If a CO ligand is replaced by a ligand that does not accept back donation from the metal, the extent of back donation is increased to the CO *trans* to the entering ligand. Therefore, that CO is bonded to the metal more strongly and is more resistant to removal. As a result, when three CO ligands are replaced, the remaining three are more difficult to replace and the substitution does not continue. Because CO exerts a strong *trans* effect, loss of CO ligands occurs so that the three entering groups are all *trans* to CO.

The structure shown above for  $\text{Mo}(\text{CO})_3(\text{py})_3$  is the result of the difference in the ability of CO and py to accept back donation from the metal. In the case where good  $\pi$  acceptors are the entering ligands, all of the CO groups may be replaced as shown in case of the reaction of  $\text{Ni}(\text{CO})_4$  with  $\text{PF}_3$ . These substitution reactions show that the electron acceptor properties of CO influence the replacement reactions.

Tracer studies using isotopically labeled CO have shown that four CO ligands in  $\text{Mn}(\text{CO})_5\text{Br}$  undergo exchange with  $^{14}\text{C}$ , but the fifth (bonded *trans* to Br) does not.

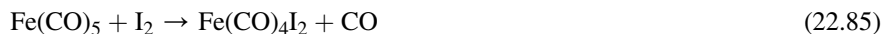


The structure of  $\text{Mn}(\text{CO})_5\text{Br}$  is



The CO *trans* to Br is held more tightly than the other four because Br does not compete effectively with CO for  $\pi$  bonding electron density donated from Mn. The other four CO groups, which are all good  $\pi$  acceptors, cause the CO groups *trans* to each other to be replaced more easily.

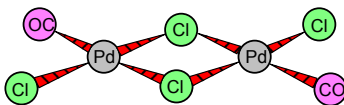
2. *Reactions with halogens.* Carbonyl halide complexes are formed by the replacement of CO with a halogen to give products that usually obey the EAN rule. For example,



A few carbonyl halides are produced directly by the reaction of a metal halide with CO.



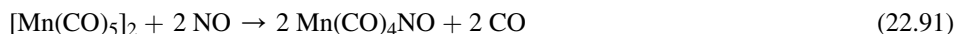
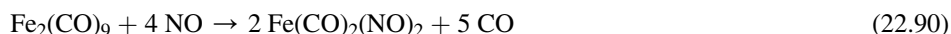
The structure of  $[\text{Pd}(\text{CO})\text{Cl}_2]_2$  involves two bridging  $\text{Cl}^-$  ions and can be shown as



3. *Reactions with NO.* Nitric oxide has one unpaired electron residing in a  $\pi^*$  orbital. When the NO molecule bonds to metals, it acts as if it donates *three* electrons. After losing one electron,



which is transferred to the metal, the resulting  $\text{NO}^+$  (the nitrosyl ion) is isoelectronic with CO and  $\text{CN}^-$ . The products containing nitric oxide and carbon monoxide, called *carbonyl nitrosyls*, are prepared by the following reactions that can be considered typical of the reactions of NO.



It is interesting to note that the products of these reactions obey the EAN rule. For example, in  $\text{Co}(\text{CO})_3\text{NO}$  the Co has 27 electrons, and it gains 3 from NO and 6 from the 3 CO ligands.

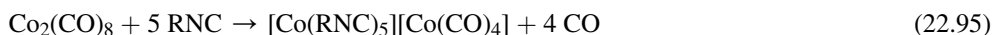
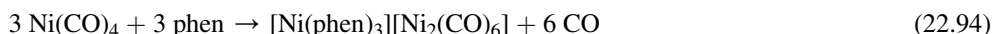
4. *Disproportionation reactions.* Metal carbonyls undergo disproportionation reactions in the presence of other potential ligands. For example, in the presence of amines,  $\text{Fe}(\text{CO})_5$  reacts as follows.



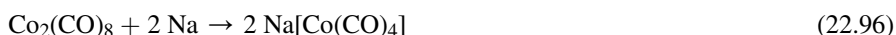
The reaction of  $\text{Co}_2(\text{CO})_8$  with  $\text{NH}_3$  is similar.



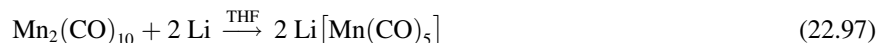
Each of these cases can be considered as a disproportionation that produces a positive metal ion and a metal ion in a negative oxidation state. The carbonyl ligands will be bound to the softer metal species, the metal anion, and the nitrogen donor ligands (hard Lewis bases) will be bound to the harder metal species, the cation. These disproportionation reactions are quite useful in the preparation of a variety of carbonylate complexes. For example, the  $[\text{Ni}_2(\text{CO})_6]^{2-}$  and  $[\text{Co}(\text{CO})_4]^-$  ions can be prepared by the reactions



5. *Carbonylate anions.* Some common carbonylate ions that obey the EAN rule are  $\text{Co}(\text{CO})_4^-$ ,  $\text{Mn}(\text{CO})_5^-$ ,  $\text{V}(\text{CO})_6^-$ , and  $\text{Fe}(\text{CO})_4^{2-}$ . The preparation of these ions is carried out by the reaction of a metal carbonyl with a substance that loses electrons easily (a strong reducing agent). Among the strongest reducing agents are the active metals such as those in Group IA. Therefore, the reactions of metal carbonyls with alkali metals would be expected to produce carbonylate ions. The reaction of  $\text{Co}_2(\text{CO})_8$  with Na carried out in liquid ammonia at  $-75^\circ\text{C}$  is one such reaction.

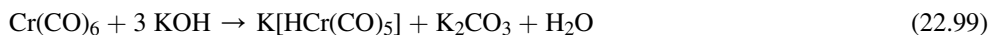


A similar reaction occurs between  $\text{Mn}_2(\text{CO})_{10}$  and Li.



Note that the anions  $\text{Co}(\text{CO})_4^-$  and  $\text{Mn}(\text{CO})_5^-$  obey the EAN rule.

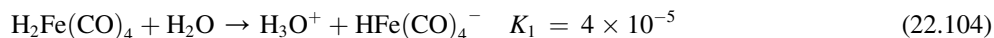
The reactions of metal carbonyls with strong bases also lead to the formation of carbonylate anions, and the following equations represent typical reactions of that type.



6. *Carbonyl hydrides.* Carbonyl hydrides are obtained by acidifying solutions containing the corresponding carbonylate anion or by the reaction of metal carbonyls with hydrogen as shown in the following reactions



A solution containing  $[\text{HFe}(\text{CO})_4]^-$  can be acidified to give  $\text{H}_2\text{Fe}(\text{CO})_4$ , which is a weak acid. The formation of the acid and its dissociation are illustrated by the following equations



Solids containing the  $\text{Fe}(\text{CO})_4^{2-}$  ion can be obtained if the cation is a large, soft one like  $\text{Hg}^{2+}$ ,  $\text{Pb}^{2+}$ , or  $\text{Ba}^{2+}$  (see Chapter 6).

## 22.6 METAL–OLEFIN COMPLEXES

It has long been known that electrons in  $\pi$  orbitals of olefins could be donated to metals to form coordinate bonds. Because of the favorable match of hard–soft character, metal–olefin complexes usually involve metals in low oxidation states. Other ligands may also be present, but many complexes are known that contain only the metal and organic ligands. This type of chemistry has its origin in the work of W. C. Zeise who prepared  $\text{K}[\text{Pt}(\text{C}_2\text{H}_4)\text{Cl}_3]$  and  $[\text{PtCl}_2(\text{C}_2\text{H}_4)]_2$  (a bridged compound) in about 1825, and it is an active area of research today. Before discussing the preparation and reactions of metal–olefin complexes, a description of the relevant structures and bonding concepts will be presented.

### 22.6.1 Structure and Bonding

In the anion of Zeise's salt, the ethylene molecule is perpendicular to the plane containing the  $\text{Pt}^{2+}$  and the three  $\text{Cl}^-$  ions as shown in Figure 22.9(a). A  $\pi$  orbital of the  $\text{C}_2\text{H}_4$  is the electron donor giving a  $\sigma$  bond to the metal as shown in Figure 22.9(b). However, the  $\pi^*$  orbitals in  $\text{C}_2\text{H}_4$  are empty, and they can accept electron density back donated from the metal, so there is a significant amount of multiple bond character to the metal–ligand bond. This situation is shown in Figure 22.10.

In  $[\text{PtCl}_2(\text{C}_2\text{H}_4)]_2$ , the double bonds of the olefin molecules lie perpendicular to the plane that contains the Pt and Cl groups. In this bridged compound, two chloride ions function as the bridging groups and the ethylene molecules are *trans* to each other as shown in Figure 22.11.

The EAN rule is useful for interpreting how ligands with more than one double bond are attached to the metal. Essentially, each double bond that is coordinated to the metal functions as an electron pair donor. Among the most

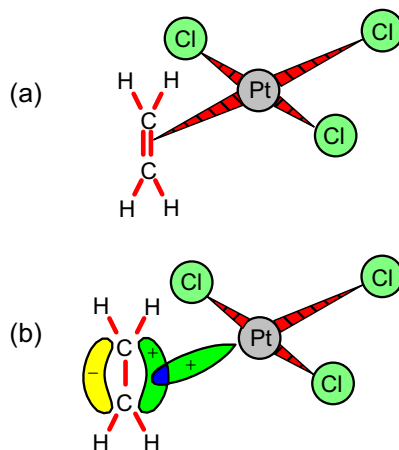


FIGURE 22.9 (a) The structure of  $[\text{PtCl}_3(\text{C}_2\text{H}_4)]^-$  and (b) the Pt–olefin  $\sigma$  bond.

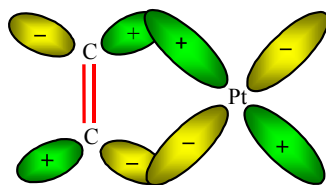


FIGURE 22.10 The overlap of the  $\pi^*$  orbital on ethylene with a  $d$  orbital on  $\text{Pt}^{2+}$ .

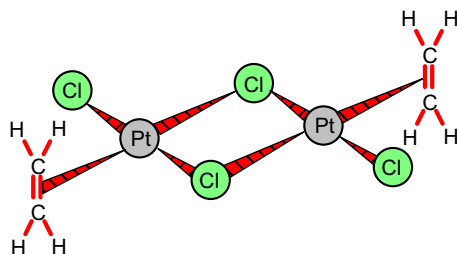


FIGURE 22.11 The structure of  $[\text{PtCl}_2(\text{C}_2\text{H}_4)_2]$ .

interesting olefin complexes are those that also contain CO as ligands. Metal–olefin complexes are frequently prepared from metal carbonyls, which undergo substitution reactions.

One ligand that possesses multiple bonding sites is cyclohepta-1,3,5-triene (cht) that contains three double bonds.



The cht ligand bonds in different ways in the complexes  $\text{Ni}(\text{CO})_3(\text{cht})$ ,  $\text{Fe}(\text{CO})_3(\text{cht})$ , and  $\text{Cr}(\text{CO})_3(\text{cht})$ . Nickel has 28 electrons and gains 6 from the 3 CO ligands. Thus, it needs only two additional electrons from cht to obey the EAN rule, and only one double bond in cht will be coordinated in  $\text{Ni}(\text{CO})_3(\text{cht})$ .

In the other complexes, six electrons also come from the three CO molecules. Because iron needs to gain a total of 10 electrons to obey the EAN rule, the cht will coordinate using 2 double bonds. In the case of the chromium complex, cht will be coordinated to all 3 double bonds in order to give a total of 36 electrons around Cr. Structures of these complexes are shown in Figure 22.12.

The bonding ability of a ligand (known as its *hapticity*) is indicated by the term *hapto* (which we will designate as  $h$ , but the symbol  $\eta$  is also used). When an organic group is bound to a metal by only one carbon atom by a  $\sigma$  bond, the bonding is referred to as *monohapto*, and it is designated as  $h^1$ . When a  $\pi$  bond in ethylene functions as the electron donor, *both* carbon atoms are considered to be bonded to the metal and the bond is designated as  $h^2$ . The bonding of cht to Ni described above is  $h^2$  because the double bond connects two carbon atoms and only one double bond functions as an electron pair donor. When two double bonds spanning four carbon atoms are functioning as electron pair donors the bonding is designated as  $h^4$ . If all three double bonds are electron pair donors, the bonding of cht is  $h^6$ . The

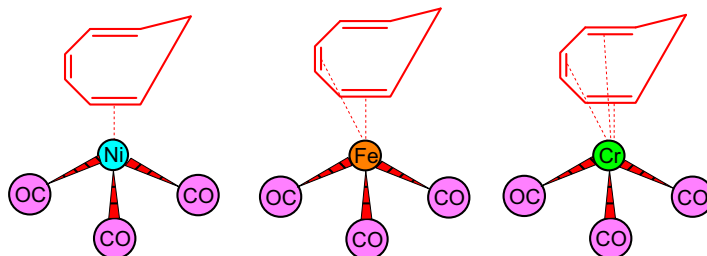
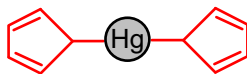


FIGURE 22.12 Bonding in complexes containing cyclohepta-1,3,5-triene.

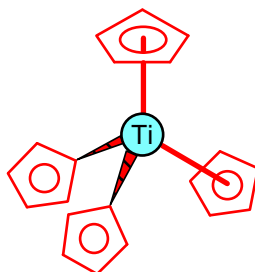


hapticity of the ligand is indicated in the formula and name of the complex. For example,  $[h^2\text{-chtNi}(\text{CO})_3]$  has the name tricarbonyldihaptocycloheptatrienenickel(0). The iron complex is written as  $[h^4\text{-chtFe}(\text{CO})_3]$  and is named tricarbonyltetrahaptocycloheptatrieneiron(0).

In ferrocene, the cyclopentadiene is bonded by the complete  $\pi$  system to the iron and, therefore, it is bonded as  $h^5\text{-C}_5\text{H}_5$ . In other cases,  $\text{C}_5\text{H}_5^-$  can bind to metals using a localized  $\sigma$  bond ( $h^1$ ). A compound of this type is  $\text{Hg}(\text{C}_5\text{H}_5)_2$ , which has the structure

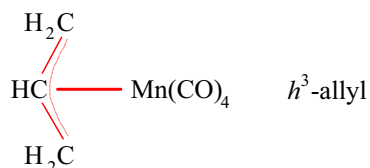
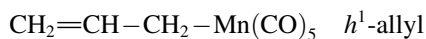


Another interesting compound that shows the different bonding abilities of cyclopentadiene is  $\text{Ti}(\text{C}_5\text{H}_5)_4$ . In this compound, two of the cyclopentadienyl ions are coordinated as  $h^5$  and the other two are bound through only one carbon atom ( $h^1$ ) in  $\sigma$  bonds to the metal. This compound has the structure



Therefore, the formula for the compound is written as  $(h^5\text{-C}_5\text{H}_5)_2(h^1\text{-C}_5\text{H}_5)_2\text{Ti}$  to show the different bonding modes of the cyclopentadienyl rings.

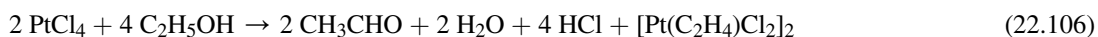
Other organic ligands can also be bound in more than one way. For example, the allyl group can be bound in a  $\sigma$  bond to one carbon atom ( $h^1$ ) or as a  $\pi$  donor encompassing all three of the carbon atoms ( $h^3$ ). These bonding modes of the allyl ligand are shown below



### 22.6.2 Preparation of Metal–Olefin Complexes

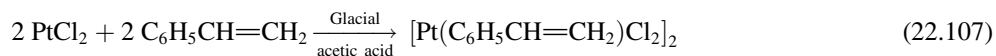
As in the case of metal carbonyls, there are several general methods for preparing metal–olefin complexes. In this section, a few of the more general methods will be described.

1. *Reaction of an alcohol with a metal halide.* In this method that can be used to obtain Zeise's salt,  $\text{K}[\text{Pt}(\text{C}_2\text{H}_4)\text{Cl}_2]$ , the dimer  $[\text{Pt}(\text{C}_2\text{H}_4)\text{Cl}_2]_2$  is obtained first and the potassium salt is obtained by treating a concentrated solution of the dimer with KCl.



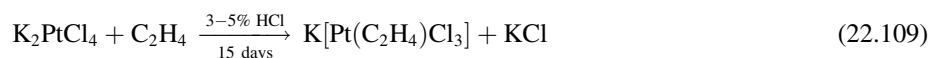
Analogous reactions can be used to prepare numerous other complexes.

2. *Reaction of a metal halide with an olefin in a nonaqueous solvent.* Some reactions of this type are the following.

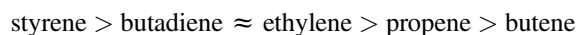


In the latter of these reactions, butadiene utilizes both double bonds by bonding to two different metal ions to give a bridged complex.

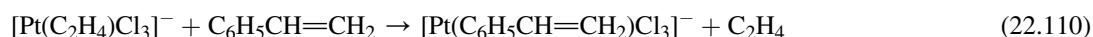
3. *Reaction of a gaseous olefin with a solution of a metal halide.* The classic synthesis of Zeise's salt,  $\text{K}[\text{Pt}(\text{C}_2\text{H}_4)\text{Cl}_3]$  is an example of this type of reaction.



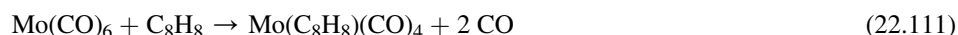
4. *Olefin substitution reactions.* Some olefins form more stable complexes than others. Therefore, it is possible to carry out substitution reactions in which one olefin replaces another. For several common olefins, the order of stability of complexes analogous to Zeise's salt is



The following is a typical replacement reaction of this type.

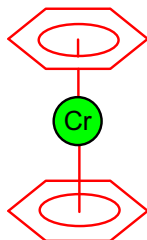


5. *Reactions of a metal carbonyl with an olefin.* In the reaction of  $\text{Mo}(\text{CO})_6$  with cyclooctatetraene, the olefin replaces two CO ligands.

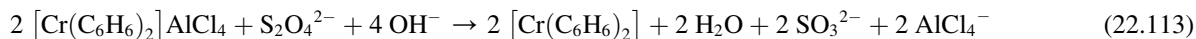


## 22.7 COMPLEXES OF BENZENE AND RELATED AROMATICS

We have considered compounds of  $\text{C}_5\text{H}_5^-$  as though the ligand is a donor of six electrons, but benzene can also function as a six-electron donor. Therefore, for a metal that has 24 electrons the addition of 2 benzene molecules would raise the total to 36, which is exactly the case if the metal is  $\text{Cr}^0$ . Thus,  $\text{Cr}(\text{C}_6\text{H}_6)_2$  obeys the EAN rule, and its structure is



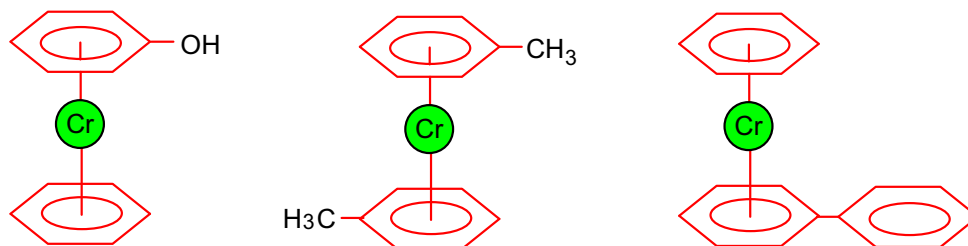
This compound has been prepared by several means including the following.



Bis(benzene)chromium(0) is rather easily oxidized, but mixed complexes can be obtained by means of substitution reactions. For example, benzene will replace three CO ligands in chromium hexacarbonyl.

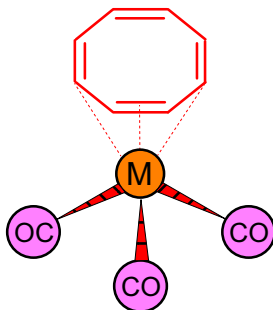


Substituted benzenes and other aromatic molecules form complexes with  $\text{Cr}(0)$  such as the following



Although the chromium compound is the best known of the benzene complexes, other metals form similar complexes with benzene and its derivatives.

Another aromatic molecule containing six  $\pi$  electrons is  $C_7H_7^+$ , the tropylium ion, derived from cycloheptatriene. This positive ion forms fewer complexes than does benzene, and they are less thoroughly studied. A molecule that has 10 electrons and has an aromatic structure is the cyclooctatetraenyl ion,  $C_8H_8^{2-}$ . Some sandwich compounds containing this ligand are known as well as complexes of the type



where  $M = Cr, Mo,$  or  $W$ . As described above, these complexes obey the 18-electron rule.

Being a hybrid field between organic and inorganic chemistry, the chemistry of metal–olefin complexes and organometallic chemistry have developed at a rapid rate. There is no doubt that this type of chemistry will be the focus of a great deal of research for some time to come.

## REFERENCES FOR FURTHER READING

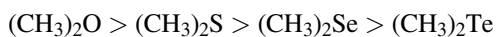
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## PROBLEMS

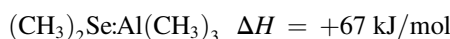
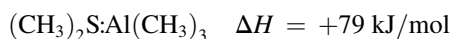
- Suppose a mixed metal carbonyl contains one Mn atom and one Co atom. How many CO molecules would be present in the stable compound? What would be its structure?
- Which of the following would be most stable? Explain your answer.
  - $Fe(CO)_2(NO)_3$
  - $Fe(CO)_6$
  - $Fe(CO)_3$
  - $Fe(CO)_2(NO)_2$
  - $Fe(NO)_5$
- Which of the following would be most stable? Explain your answer.
  - $Fe(CO)_4NO$
  - $Co(CO)_3NO$
  - $Ni(CO)_3NO$
  - $Mn(CO)_6$
  - $Fe(CO)_3(NO)_2$
- How is benzene bound to the metal in  $Cr(CO)_3(C_6H_6)$ ?

5. Describe the structure and bonding in the following compounds where  $C_4H_6$  is butadiene and  $C_6H_6$  is benzene.
- $Ni(C_4H_6)(CO)_2$
  - $Fe(C_4H_6)(CO)_4$
  - $Cr(C_6H_6)(CO)(C_4H_6)$
  - $Co(C_4H_6)(CO)_2(NO)$
6. What is the structure (show all bonds clearly) of  $Co(CO)_2(NO)(cht)$  where cht is cycloheptatriene.
7. Draw the structure of  $Mn(CO)_2(NO)(cht)$  and explain how the molecule does or does not obey the 18-electron rule.
8. Show structures (show all bonds clearly) for the following where  $C_8H_8$  is cyclooctatetraene.
- $Fe(CO)_3(C_8H_8)$
  - $Cr(CO)_3(C_8H_8)$
  - $Co(CO)(NO)(C_8H_8)$
  - $Fe(CO)_3(C_8H_8)Fe(CO)_3$
9. Show structures (show all bonds clearly) for the following where  $C_8H_8$  is cyclooctatetraene.
- $Ni(CO)_2(C_8H_8)$
  - $Ni(NO)_2(C_8H_8)$
  - $Cr(CO)_4(C_8H_8)$
  - $Ni(CO)_3(C_8H_8)Cr(CO)_4$
10. Explain why  $Co(CO)_4$  is not a stable compound but  $Co_2(CO)_8$  is. Draw structures for materials having the formula  $Co_2(CO)_8$ . How could you determine the structures for these materials?
11. Predict the products of the reactions indicated. More than one reaction may be possible.
- $Fe_2(CO)_9 + NO \rightarrow$
  - $Mn_2(CO)_{10} + NO \rightarrow$
  - $V(CO)_6 + NO \rightarrow$
  - $Cr(CO)_6 + NO \rightarrow$
12. Complete and balance the following reactions:
- $LiGe(C_6H_5)_3 + (C_2H_5)_3SiCl \rightarrow$
  - $Li + C_4H_9Cl \rightarrow$
  - $(CH_3)_3SnBr + C_6H_5MgBr \rightarrow$
  - $TiOH + C_5H_6 \rightarrow$
  - $Sn(CH_3)_4 + Br_2 \rightarrow$
13. Complete and balance the following reactions:
- $B(CH_3)_3 + O_2 \rightarrow$
  - $LiC_4H_9 + Br_2 \rightarrow$
  - $Zn(C_2H_5)_2 + O_2 \rightarrow$
  - $Hg(C_2H_5)_2 + HgBr_2 \rightarrow$
  - $Zn(C_2H_5)_2 + NH_3 \rightarrow$
14. Complete and balance the following reactions:
- $Ga(CH_3)_3 + I_2 \rightarrow$
  - $Hg(CH_2C_6H_5)_2 + Na \rightarrow$
  - $[Al(CH_3)_3]_2 + (C_2H_5)_3N \rightarrow$
  - $NaC_2H_5 + C_6H_5Cl \rightarrow$
  - $Zn(CH_3)_2 + GeCl_4 \rightarrow$
15. On the basis of the hard–soft interaction principle, explain why the following reactions take place.
- $Sb(C_6H_5)_3 + As \rightarrow Sb + As(C_6H_5)_3$
  - $As(C_6H_5)_3 + P \rightarrow (C_6H_5)_3P + As$
16. Explain why  $Al(CH_3)_3$  forms a more stable complex with  $(CH_3)_3N$  than it does with  $(CH_3)_3P$ .
17. When weak Lewis bases that form complexes with trimethyl gallium are added to trimethylaluminum, no reaction occurs. What is the origin of this difference?

18. Explain the fact that toward aluminum alkyls, the ability to form complexes varies in the order



19. Explain the difference in the heats of decomposition of adducts shown below.



20. Draw the structure of the product that results.
- Mo(CO)<sub>6</sub> reacting with excess pyridine
  - The reaction of cobalt with CO at high temperature and pressure
  - Mn<sub>2</sub>(CO)<sub>10</sub> reacting with NO
  - Fe(CO)<sub>5</sub> reacting with cycloheptatriene
21. During a study of Cr(CO)<sub>6</sub>, Ni(CO)<sub>4</sub>, and Cr(NH<sub>3</sub>)<sub>3</sub>(CO)<sub>3</sub> by IR spectroscopy, three spectra were obtained showing CO stretching bands at 1900 cm<sup>-1</sup>, 1980 cm<sup>-1</sup>, and 2060 cm<sup>-1</sup>. Identify which peak corresponds to each compound and explain your answer.
22. For a complex M(CO)<sub>5</sub>L, two bands are observed in the region 1900–2200 cm<sup>-1</sup>. Explain what this observation means. Suppose L can be NH<sub>3</sub> or PH<sub>3</sub>. When L is changed from NH<sub>3</sub> to PH<sub>3</sub>, one band is shifted in position. Will it be shifted to higher or lower wave numbers? Explain.
23. Infrared spectra of Ni(CO)<sub>4</sub>, CO(g), Fe(CO)<sub>4</sub><sup>2-</sup>, and Co(CO)<sub>4</sub><sup>-</sup> have bands at 1790, 1890, 2143, and 2060 cm<sup>-1</sup>. Match the bands to the appropriate species and explain your reasoning.
24. What is the name for (h<sup>5</sup>-C<sub>5</sub>H<sub>5</sub>)<sub>2</sub>(h<sup>1</sup>-C<sub>5</sub>H<sub>5</sub>)<sub>2</sub>Ti?
25. For the anion [C<sub>2</sub>H<sub>4</sub>PtCl<sub>3</sub>]<sup>-</sup>, how would you expect the stretching frequency for the C=C bond to compare to that in gaseous ethylene? Explain your answer.
26. Does the compound (h<sup>5</sup>-C<sub>5</sub>H<sub>5</sub>)Fe(CO)<sub>2</sub>Cl obey the 18-electron rule? Explain.
27. Experimental evidence suggests that there is some back donation from platinum to ethylene in Zeise's salt. Make an appropriate sketch to show the bonding and explain how this could occur.
28. Explain how the existence of isomers of substituted ferrocene molecules such as those containing two alkyl groups would be different if rotation of the rings were not possible.

# Inorganic Substances in Biochemical Applications

Although the image of inorganic chemistry is associated with the production of sulfuric acid, ammonia, and lime, there is another side to the field. Many living systems require “inorganic” elements in order to function even though the vast majority of living matter is composed primarily of carbon, hydrogen, oxygen, nitrogen, sulfur, and phosphorus. Because the preponderance of living matter is made of organic matter, it is to be expected that the most abundant elements would be those listed. Moreover, much of the hydrogen and oxygen content of organic matter is in the form of water. In addition to these constituents several metals are important among which are iron, zinc, nickel, manganese, cobalt, copper, and vanadium. Other metallic elements found as ions are  $\text{Na}^+$ ,  $\text{K}^+$ ,  $\text{Mg}^{2+}$  and  $\text{Ca}^{2+}$ . For example, chlorophyll contains a complex containing magnesium, and the transport of oxygen in blood is accomplished by hemoglobin that is an iron complex.

Traces of metals such as copper, zinc, and chromium are also required by living systems. The role of compounds that contain elements other than carbon, hydrogen, and oxygen is a vast and rapidly growing field. Consequently, this chapter will present only a brief overview of the field as it relates the importance of several elements in naturally occurring systems and medicinal compounds. In that way, the breadth of the range of topics will be illustrated. Just as a few highlights can show the important plays of a game, brief looks at some of the inorganic materials in biological systems can illustrate the nature of the field.

Throughout this chapter, numerous references will be made to the principles of molecular structure and molecular interactions discussed in earlier chapters. For more biological aspects of the topics discussed, the reader should consult the references listed at the end of this chapter.

## 23.1 THERAPEUTIC ASPECTS OF INORGANIC SUBSTANCES

Although the role of certain metal ions in the action of enzymes is extremely important, numerous inorganic substances are useful in much less exotic ways. In this section, some of the situations will be described in which compounds containing inorganic constituents are employed in these “medicinal” ways. The small number of instances described is meant to be indicative of the scope of such uses, but anything approaching a comprehensive treatment is impractical in a book such as this. It is hoped that the topics presented will serve to not only show the wide range of inorganic chemistry, but also to initiate an interest on the part of the reader to study this field more fully. Table 23.1 lists numerous inorganic substances that have medicinal applications.

Although many of the substances shown in Table 23.1 are familiar and used in products that are commonplace, that does not mean that they present no hazard. Even products such as toothpaste carry warnings about preventing ingestion by young children because fluorine-containing compounds are toxic.

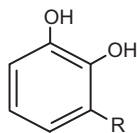
### 23.1.1 Metal Compounds in Treatment of Poison Ivy

Since ancient times, substances that are wholly or in part inorganic in composition have been used for medicinal purposes. Some of these uses have been prevalent in folk medicine for centuries for common ailments. For example, decades ago a treatment for poison ivy (*Toxicodendron radicans*, a member of the *Toxicodendron* group that is also known as *Rhus radicans*) was to obtain some “blacksmith’s water” and apply it to the affected area. Blacksmith’s water was that in which the smith dipped hot iron objects that were being hammered into shape. Although iron does not react with water at ordinary temperatures, it does at high temperature (see Section 7.1). Therefore, over time the water became a solution containing  $\text{Fe}^{2+}$ ,  $\text{Fe}^{3+}$ , or a mixture of these ions. The lead author had a great uncle who was a blacksmith and he followed such a treatment on numerous occasions in his youth. Interestingly, *calamine*, a mixture that contains  $\text{ZnO}$  and  $\text{Fe}_2\text{O}_3$ , is a widely used remedy in treating poison ivy. Another suggested treatment or preventative is a solution of ferric chloride in alcohol (National Safety Council Data Sheet 304, June 2006).

**TABLE 23.1** Some Common Inorganic Chemicals That Have Been Used in Medicines

Compound	Use
Aluminum hydroxide	Antacid
Barium sulfate	X-ray contrast agent
Bismuth subsalicylate	Indigestion (Pepto-Bismol™)
Boric acid	Eye wash, disinfectant
Calcium carbonate	Antacids, calcium supplement
Copper(II) sulfate	Copper supplement
Ferric chloride	Treatment of poison ivy
Gadolinium compounds	MRI contrast agents
Hydrogen peroxide	Disinfectant
Iodine	Antiseptic
Iodized salt	Iodine supplement to prevent goiter
Iron(II) sulfate	Iron supplement
Lithium carbonate	Treatment of bipolar disorder
Magnesium hydroxide	Antacid (Milk of Magnesia™)
Magnesium oxide	Antacid, laxative
Magnesium sulfate	Epsom salts
Mercurochrome	Mercury-containing compound (formerly an antiseptic)
Nitrous oxide	Dental anesthetic
Selenium sulfide	Dandruff shampoo
Sodium bicarbonate	Antacid
Sodium borate	Eye wash, mouth rinse
Sodium bromide	Formerly used as a tranquilizer, painkiller (toxic!)
Sodium fluoride	Fluoride source in toothpastes
Sodium <sup>131</sup> Iodide	Diagnosis of thyroid problems
Sodium monofluorophosphate	Fluoride source in toothpaste
Stannous fluoride	Fluoride source in toothpaste
Titanium dioxide	Sunscreen
Zinc oxide	Skin rash cream, sunscreen

The toxic substance in poison ivy is urushiol, a mixture of derivatives of catechol, 1,2-dihydroxybenzene,  $C_6H_4(OH)_2$ , having long alkyl groups containing 15–17 carbon atoms in a position *ortho* to one of the hydroxyl groups.

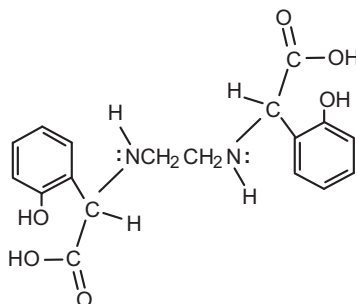


This material can be absorbed through the skin causing an itching rash. Compounds of this type are acidic and easily form the anion or dianion by loss of protons. The dianion forms stable chelate complexes with  $Fe^{3+}$  so it is possible that the historic use of blacksmith's water to treat poison ivy has some validity. Studies have also been conducted in which  $MnSO_4$  was administered to relieve the symptoms of poison ivy (Terezakis, et al. 1995). The results obtained were consistent with the  $Mn^{2+}$  ion forming stable complexes with the urushiol. In this somewhat arcane way, the principles of coordination chemistry are applicable to medicinal science.

### 23.1.2 Iron Deficiency in Trees

As will be shown later in this chapter, iron is required for the transport of oxygen by blood in humans. However, iron is also needed by other biological systems. For example, several species of trees, among them the oaks, require iron or else they develop a condition known as *interveinal chlorosis*. When there is a deficiency of iron in the trees, the leaves have a very pale green or yellowish appearance. This state is the result of inadequate uptake of iron from the soil, usually brought about by the alkalinity of the soil. The hydroxides of transition metals are insoluble and are easily converted to the oxides, which are also insoluble.

Treatment of the trees or the soil can correct the condition and sometimes a tree may be given injections of ferric ammonium citrate or ferrous sulfate. In other cases, sulfuric acid may be added to the soil around the tree, which renders the iron in a more soluble form. To make iron available to the trees, iron in the form of stable but soluble chelates can be added. One of the most common chelating ligands is known as EDDHA, an abbreviation for 2-[2-[[2-Hydroxy-1-(2-hydroxyphenyl)-2-oxoethyl]amino]ethylamino]-2-(2-hydroxyphenyl)acetic acid. Like EDTA, EDDHA is capable of binding to metal ions at numerous sites. The structure of this chelating agent can be shown as



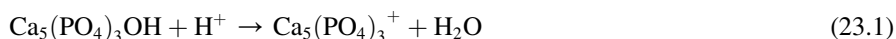
The ability of ligands such as EDTA and EDDHA to form stable complexes with transition metals is also utilized in other ways. For example, salad dressings often have in the list of ingredients a statement to the effect that EDTA is added to retard spoilage. The ligand is effective in this way because it binds metal ions that otherwise catalyze oxidation reactions leading to spoilage. In other words, the metal ions are *sequestered*, leading to the term *sequestering agents* as applied to strongly binding chelating agents.

### 23.1.3 Fluorides

Well over a century ago, sodium fluoride was used as an insecticide and also as a rat poison. These uses are not without some risk because of the toxicity of NaF. In spite of the fact that in sizable amounts NaF is toxic, in very low concentrations it has found application in some toothpastes. Moreover, some brands of toothpaste that do not contain NaF make use of sodium monofluorophosphate as the fluorinating agent. Another fluorine compound that has been used in toothpaste is stannous fluoride, which was marketed under the name Fluoristan<sup>®</sup>. This use of SnF<sub>2</sub> was subsequently discontinued and sodium monofluorophosphate, Na<sub>2</sub>PO<sub>3</sub>F, became the fluorinating agent of choice. It is sometimes referred by the trade name Fluoristat<sup>®</sup>.

The mineral *apatite* is a form of calcium phosphate. However, numerous derivatives also exist in which groups such as OH, F, or Cl are present. As a result, the formula is sometimes shown as Ca<sub>10</sub>(PO<sub>4</sub>)<sub>6</sub>(OH, F, Cl)<sub>2</sub>. This formula indicates that there are two negative ions but whether they are OH<sup>-</sup>, F<sup>-</sup>, or Cl<sup>-</sup> is not specified, and there may in fact be a mixture of these ions as long as there are two. The simplest formula would contain five Ca<sup>2+</sup> and three PO<sub>4</sub><sup>3-</sup> ions if only one of the univalent anions were present. In those cases, the compounds are *hydroxyapatite*, Ca<sub>5</sub>(PO<sub>4</sub>)<sub>3</sub>OH, *fluoroapatite*, Ca<sub>5</sub>(PO<sub>4</sub>)<sub>3</sub>F, and *chloroapatite*, Ca<sub>5</sub>(PO<sub>4</sub>)<sub>3</sub>Cl.

Teeth have a surface that contains *hydroxyapatite*. The reaction leading to tooth decay can be shown as



In the presence of F<sup>-</sup>, the surface is converted into Ca<sub>5</sub>(PO<sub>4</sub>)<sub>3</sub>F, which does not undergo reaction with acids as readily as does hydroxyapatite. When used as an ingredient in toothpaste, sodium monofluorophosphate constitutes 0.76%. Both sodium monofluorophosphate and sodium fluoride continue to be used in toothpastes.

### 23.1.4 Titanium Dioxide

Titanium dioxide occurs naturally in the minerals *rutile anastase*, and *ilmenite*. The compound has many uses, most of which are related to its brilliant white color. It is used as an ingredient in paints in which it is the opaque covering agent even though the paint may have color as the result of added pigments. The quantity of TiO<sub>2</sub> used in this way is enormous.



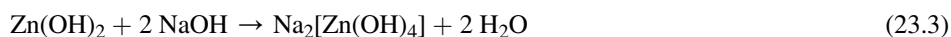
Titanium dioxide is also used in many other applications that include preparation of glazes, electronics components, and papers. Among its other uses are in cosmetics, some toothpaste as a brightening agent, and more recently, in catalysts. Some sunscreens contain  $\text{TiO}_2$ , which absorbs UV light. Certain brands of powdered sugar donuts have included  $\text{TiO}_2$  as an ingredient. However, some producers of these products have discontinued this practice. Although not highly toxic when compared to compounds of heavy metals, in a finely divided particulate state  $\text{TiO}_2$  has been implicated in causing cancer. There appears to be considerable variation of the effects depending on the particle size, which is not unexpected based on the ease of mobility of nanoparticles through living systems. Considering the wide range of products in which  $\text{TiO}_2$  is used, it is not surprising to find it is an inorganic substance that is a health concern.

### 23.1.5 Zinc Oxide

Another inorganic compound that humans encounter frequently is zinc oxide. Although it is a metal oxide and most metal oxides are the anhydrides of bases,  $\text{ZnO}$  is an amphoteric oxide. As the hydroxide in aqueous solutions, it reacts as a base toward acids such as  $\text{HCl}$ ,



and as an acid toward  $\text{NaOH}$ ,

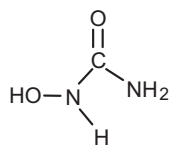


Zinc oxide is a white solid that has been used as a pigment. As with the sulfide, the two forms are *wurtzite* and *zinc blende* (see Section 4.1.3).

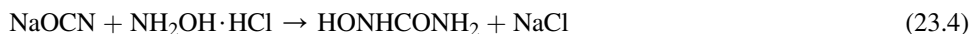
Because zinc oxide is effective at blocking ultraviolet radiation, it is widely used in sunscreens, suntan lotions, and many forms of cosmetics. It has been used in various ointments and salves since antiquity. A popular ointment known as *calamine* contains both  $\text{ZnO}$  and  $\text{Fe}_2\text{O}_3$ , the latter to the extent of approximately 0.5%. When used in a finely divided state,  $\text{ZnO}$  has antibacterial properties, which leads to its use in a wide variety of both pharmaceuticals and cosmetics. Zinc is the metal ion located in the reactive center of the enzyme *carbonic anhydrase* which results in it being classified as an essential trace metal. Because of this, zinc compounds are added to several foods. Zinc oxide is essentially nontoxic, but as with many other substances, vapors emitted from the materials heated to high temperature or as finely divided powders constitute a hazard.

### 23.1.6 Hydroxyurea

Although perhaps not strictly of an “inorganic” nature, another compound that has important pharmaceutical uses is hydroxyurea, which is also known as hydroxycarbamate. Carbamic acid is  $\text{H}_2\text{NCOOH}$  so replacing  $\text{OH}$  in the carboxyl group with an  $\text{NH}_2$  group results hydroxyurea that has the structure



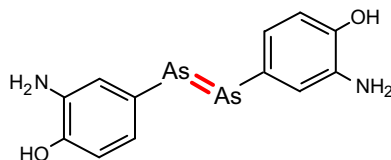
Hydroxyurea is produced by the reaction of sodium cyanate with hydroxylamine hydrochloride under basic conditions.



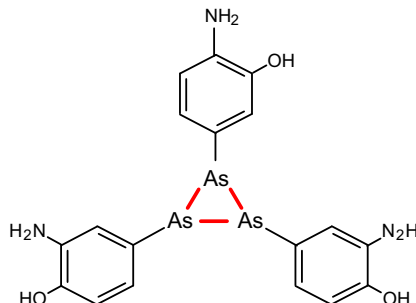
This reaction is similar to that by which cyanates are converted to urea, which was influential in eliminating the demarcation between organic and inorganic chemistry by Wöhler in 1828. Hydroxyurea *in vivo* acts by inhibiting the growth of cancer cells which gives rise to its being used in the treatment of several types of cancer, sickle cell anemia, psoriasis, and AIDS. Its primary use is in the treatment of certain types of leukemia.

### 23.1.7 Arsenic Compounds

Another element that has figured prominently in medicinal uses is arsenic. An organoarsenic compound known as *arsphenamine* or *salvarsan* was the original chemotherapeutic treatment for syphilis. The compound was long believed to be the dihydrochloride of the molecule having the structure



However, recent structural determinations have shown that there are rings containing three and five arsenic atoms present rather than two As atoms with a double bond between them (Lloyd, et al., 2005; Yarnell, 2005).

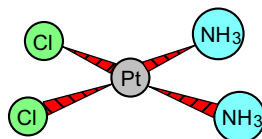


Although it will not be shown, the other compound has as its essential feature a five-membered arsenic ring in the center. More recently, antibiotics have been found that are more effective than are the arsenic compounds, but nevertheless, arspenamine and its derivatives played an important role in medicine.

For many years, a wide range of arsenic compounds, both organic and inorganic, were utilized in pesticides, insecticides, and herbicides. Other applications of arsenic compounds included its use as wood preservative and in sprays for fruit trees. Most such uses have been discontinued. Even the use of salvarsan has largely been discontinued because more effective drugs.

### 23.1.8 Cisplatin and Other Platinum Complexes

Arguably the most important inorganic compound used as a pharmaceutical is the platinum complex  $cis\text{-}[\text{Pt}(\text{NH}_3)_2\text{Cl}_2]$ , which is also known as cisplatin.



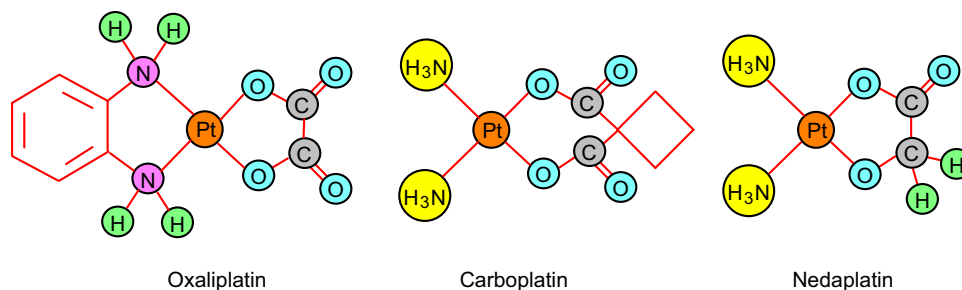
As would be expected for a complex containing a second row transition metal having a  $d^8$  configuration, the complex is square planar. The compound was first prepared in 1845 by Michel Peyrone, and it was approved for use in treatment of cancer in 1978. Although cisplatin is effective against certain types of cancer, the search for more effective platinum complexes has led to the preparation and testing of an incredible number of compounds. Hydrolysis of cisplatin occurs in cells to produce  $cis\text{-}[\text{Pt}(\text{NH}_3)_2(\text{H}_2\text{O})(\text{OH})]^+$  and  $cis\text{-}[\text{Pt}(\text{NH}_3)_2(\text{H}_2\text{O})\text{Cl}]^+$ , which become the actual effective form of the drug. Being a compound containing a heavy metal, cisplatin is toxic and can lead to kidney damage and other side effects. Another issue is that some types of cancer become resistant to the drug.

It is interesting to note that  $trans\text{-}[\text{Pt}(\text{NH}_3)_2\text{Cl}_2]$  is not effective against cancer cells. The reason may be that the *cis* isomer can form bridges between groups in DNA that prohibit cell replication and cause cell death. It may be that the *trans* isomer cannot form similar linkages between DNA groups.

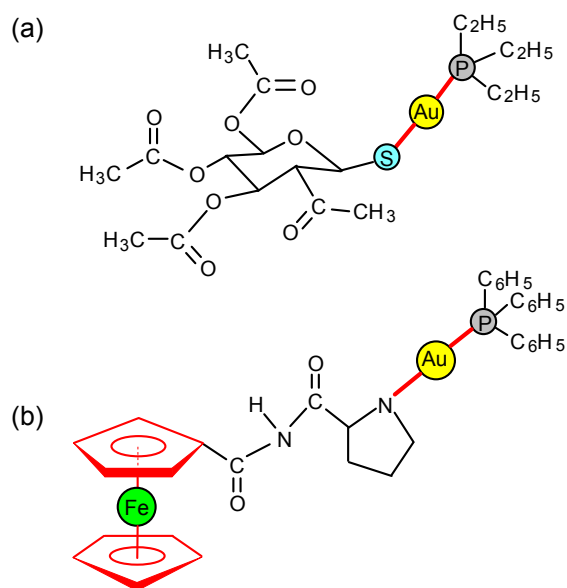
Platinum complexes are kinetically inert toward substitution (and hydrolysis) so one of the goals of producing and testing other platinum complexes is to find compounds that are equally effective but that undergo hydrolysis at different rates. Some of the important complexes that have been studied have structures that are shown in Figure 23.1. All of these compounds have been shown to be effective in treating certain types of cancer. In addition numerous complexes containing two platinum ions connected by bridging groups have also been tested.

### 23.1.9 Auranofin

Platinum is not the only precious metal that has compounds known to have therapeutic uses. Another is auranofin, a complex of gold that is used in the treatment of rheumatoid arthritis and has the structure shown in Figure 23.2(a). In addition,



**FIGURE 23.1** The structures of oxaliplatin, carboplatin, and nedaplatin. Adapted from House (2013).

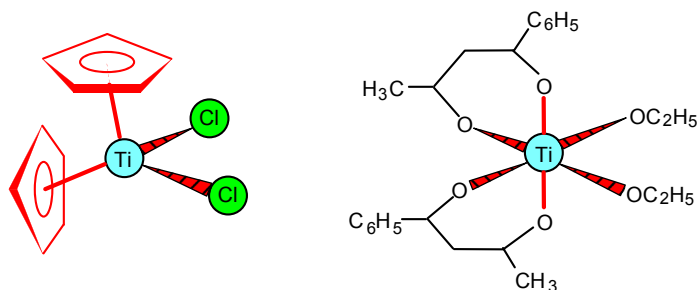


**FIGURE 23.2** The structures of two gold compounds, (a) auranofin and (b) a gold compound containing ferrocene as a ligand, used in treatment of rheumatoid arthritis. Adapted from House (2013).

numerous other gold compounds are useful not only for rheumatoid diseases but also show anticancer effectiveness. One such compound contains a ferrocene moiety and has the structure shown in Figure 23.2(b). Complexes containing osmium and ruthenium have also been shown to be effective in treating some types of cancer (see Section 23.1.11).

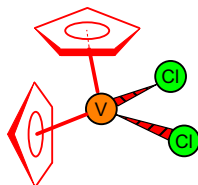
### 23.1.10 Anticancer Compounds Containing First-Row Metals

In the intense quest to develop more effective drugs for treating cancer, compounds containing many different metals have been studied. Among them are two complexes of titanium that are shown in Figure 23.3.



**FIGURE 23.3** Two titanium complexes that have been tested for use on cancer. Adapted from House (2013).

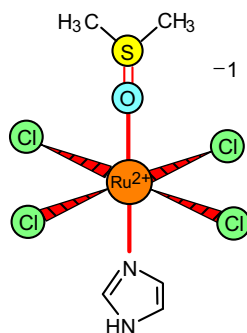
The titanium complexes have been found to be effective toward ovarian and colon cancer. However, as was discussed in Chapter 21, complexes of titanium(IV) undergo substitution readily and, therefore, hydrolyze rapidly, which is a disadvantage in cancer treatment. A large number of vanadium complexes derived from



have also been studied. Most of the derivatives contain various substituents on the cyclopentadienyl rings. The V–Cl bonds undergo hydrolysis easily, but the presence of various alkyl groups attached to the cyclopentadienyl rings reduces the solubility of the compounds and thus retards hydrolysis. In addition to these compounds, many others containing first-row transition metals have been studied as candidates for use in chemotherapy.

### 23.1.11 Ruthenium Compounds in Cancer Treatment

In addition to the numerous compounds described, complexes of ruthenium have also been found to have anticancer effectiveness. One of the most successful is the complex containing dimethylsulfoxide and an imidazole as ligands. Being an anion having a  $-1$  charge, it is utilized as the sodium salt or with a protonated imidazole as the cation. These complexes have the formulas  $\text{Na}^+[\text{trans-Ru}(\text{Cl}_4)(\text{DMSO})(\text{Im})]^-$  and  $\text{ImH}^+[\text{trans-Ru}(\text{Cl}_4)(\text{DMSO})(\text{Im})]^-$  in which Im represents an imidazole molecule. The structure of the anion can be shown as



This compound appears to be effective in treating cancer by causing death of the cells (apoptosis) rather than by inhibiting replication as in the case of cisplatin. Many other ruthenium complexes have been tested for anticancer effectiveness.

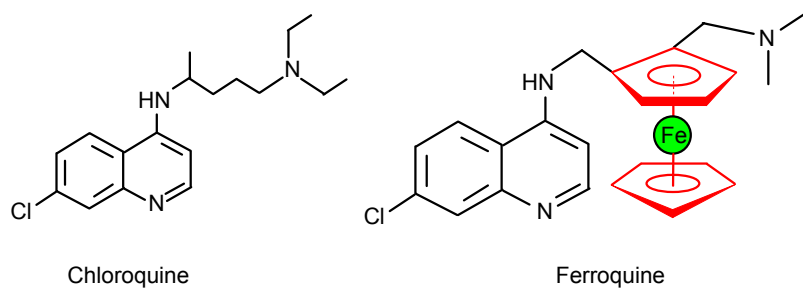
### 23.1.12 Metal Compounds in Treatment of Malaria

Another area of research in which coordination compounds containing metals figure prominently is in the treatment of malaria. On a worldwide basis, it is estimated that from 700,000 to a million people die of malaria annually. Although it was discovered in the 1600s that quinine could be used to treat victims of malaria, many newer drugs are more effective. Among them are the compound chloroquine, but a derivative that contains the ferrocene group, known as ferroquine, is even more effective, especially against parasites that are resistant to chloroquine. The structures of these compounds are shown in Figure 23.4.

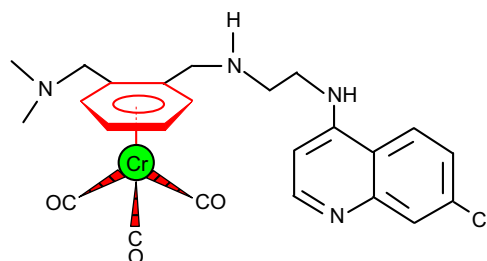
Another antimalarial drug is derived from a chromium complex that contains both three carbonyl ligands and a coordinated six-membered aromatic ring that is bonded to other groups. Its structure is shown in Figure 23.5. Several complexes that contain ruthenium are also effective in the treatment of malaria.

## 23.2 BIOCHEMICAL ASPECTS OF ENERGY CHANGES

Energy is produced by many biochemical systems whereas in others, energy is consumed as chemical reactions occur. An example of the former case is the production of energy in cells with ATP as a catalyst and an example of the latter is photosynthesis in which light is absorbed as carbohydrates are produced. The enzyme *carbonic anhydrase* requires  $\text{Zn}^{2+}$ ,



**FIGURE 23.4** The structures of chloroquine and ferroquine. Adapted from House (2013).



**FIGURE 23.5** A chromium compound that is effective in treating malaria. Adapted from Glans et al. (2011).

which is effective by causing a coordinated water molecule to undergo proton loss. A brief description of some such processes will be given with the emphasis on their “inorganic” aspects.

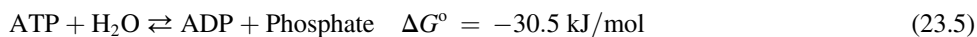
### 23.2.1 Production of Energy by ATP

Energy is produced in cells by a series of complex reactions that depend on biological catalysts known as enzymes. In some living systems, the energy production involves a cycle that is dependent on a complex material known as *adenosine triphosphate* or ATP, which functions as a coenzyme. A coenzyme is a substance that is necessary for the enzyme to serve as a catalyst for a biochemical reaction. ATP consists of a triphosphate chain that is similar to that contained in the “polyphosphoric” acids described in Section 14.8.2. Attached to the triphosphate group is the ribose sugar and attached to the sugar moiety is an adenosine ring. The structure of ATP is usually shown in the anion form with the hydrogen ions removed from the phosphate groups as shown in Figure 23.6.

A highly charged ion such as the ATP anion (or any highly charged species such as  $\text{Fe}^{3+}$ ,  $\text{PO}_4^{3-}$ , or  $\text{N}^{3-}$ ) is not normally stable in aqueous solutions. However, with the presence of multiple oxygen atoms having unshared pairs of electrons and negative charges, that portion of the ATP molecule is a good chelating agent and it binds to  $\text{Mg}^{2+}$  ions, which stabilizes the ATP in anion form. Even though this occurs, the terminal phosphate can be removed in an energetically favorable step. This occurs because of the energy of interaction of the liberated phosphate ion with the solvent ( $\text{H}_2\text{O}$ ) and the greater stability of the chelate with  $\text{Mg}^{2+}$  ions. If the  $\text{Mg}^{2+}$  ion is solvated by several water molecules, forming a chelate with the triphosphate group is favorable because of the chelate (entropy) effect as water molecules are displaced (see Section 20.4.2).

In Chapter 6, it was shown that in an acid molecule the number of oxygen atoms with no hydrogen attached affects the strength of the acid as a result of resonance. In the ATP molecule, the electrons around bridging oxygen atoms are localized to a greater extent than are those of oxygen atoms in terminal positions on the phosphate groups. Thus, there is some contribution to the favorable energy of hydrolysis as a result of increased resonance stabilization of the products compared to the ATP molecule.

The product that results when a phosphate group is removed from ATP is known as adenosine diphosphate or ADP. Thus there is an equilibrium that can be shown as



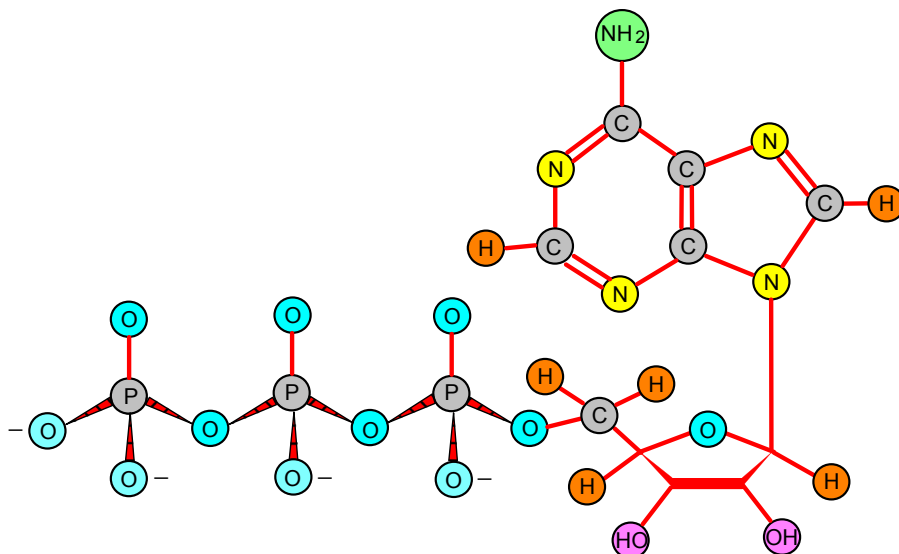


FIGURE 23.6 The structure of the adenosine triphosphate anion.

However, the ATP ion has a  $-4$  charge and the phosphate ion produced is the anion of a weak acid. Therefore, hydrolysis occurs so that the phosphate liberated can be present as  $\text{HPO}_4^{2-}$  or  $\text{H}_2\text{PO}_4^-$ , and the reactions producing these species from  $\text{PO}_4^{3-}$  are energetically favorable. As a result, the reaction illustrated in Eq. (23.5) releases 30.5 kJ/mol.

### 23.2.2 Photosynthesis

Photosynthesis is the name applied to the process whereby energy in the form of light is used as the energy necessary to produce biological materials. The processes involve the fixation of carbon and involve in simplest terms the reaction illustrated as



Not only is such a process important from the standpoint of glucose (and subsequently starch) production, but also it is the source of oxygen liberated to the atmosphere. It has been estimated that as much as 70% of the oxygen in the atmosphere is produced by plants.

Instrumental in the syntheses are the chlorophylls, of which there are many. The structure of chlorophyll *a* is shown in Figure 23.7 in which the groups labeled  $\text{R}_1$ ,  $\text{R}_2$ , and  $\text{R}_3$  are  $-\text{CH}=\text{CH}_2$  (vinyl),  $-\text{CH}_3$  (methyl), and  $-\text{CH}_2\text{CH}_3$  (ethyl), respectively. Note that in this structure the chelated metal ion in the fused ring system is  $\text{Mg}^{2+}$ . If  $\text{R}_1$  and  $\text{R}_3$  remain unchanged but  $\text{R}_2$  is changed to the aldehyde group  $-\text{CHO}$ , the result is chlorophyll *b*.

The process is much more complex than that shown in Eq. (23.6) in which glucose is the product. In most cases, the predominant product is glyceraldehyde-3-phosphate, G3P,  $\text{HOC}-\text{CH}(\text{OH})\text{CH}_2\text{OPO}_3$ . This compound is converted into glucose in the presence of an enzyme, and these glucose units are subsequently transformed enzymatically by a complex series of steps into starches, which become important raw materials as well as foodstuffs. In other situations, sucrose is the end product of reactions of G3P.

From the standpoint of inorganic chemistry, photosynthesis is the result of magnesium complexes having chromophores (color-absorbing groups). The green color of chlorophylls is due to their absorbance in the 400–500 nm (blue) and 650–700 nm (red) regions. It is generally the case that molecules containing conjugated rings systems absorb electromagnetic radiation as electrons are excited from the ground state into excited states. In some cases the transitions involve electrons being excited from a  $\pi$  bonding orbital to a  $\pi^*$  antibonding orbital (known as a  $\pi \rightarrow \pi^*$  transition) although the light absorbed is not always in the visible region of the spectrum. Pigments such as alizarin and indigo exhibit this color as a result of absorption of visible light, but benzene is colorless because the absorption is in the ultraviolet region of the spectrum. In fact, benzene exhibits a strong absorption at approximately 200 nm as a result of a  $\pi \rightarrow \pi^*$  transition. Borazine,  $\text{B}_3\text{N}_3\text{H}_6$ , contains alternating single and double bonds in the ring structure, and it also has a strong absorption in

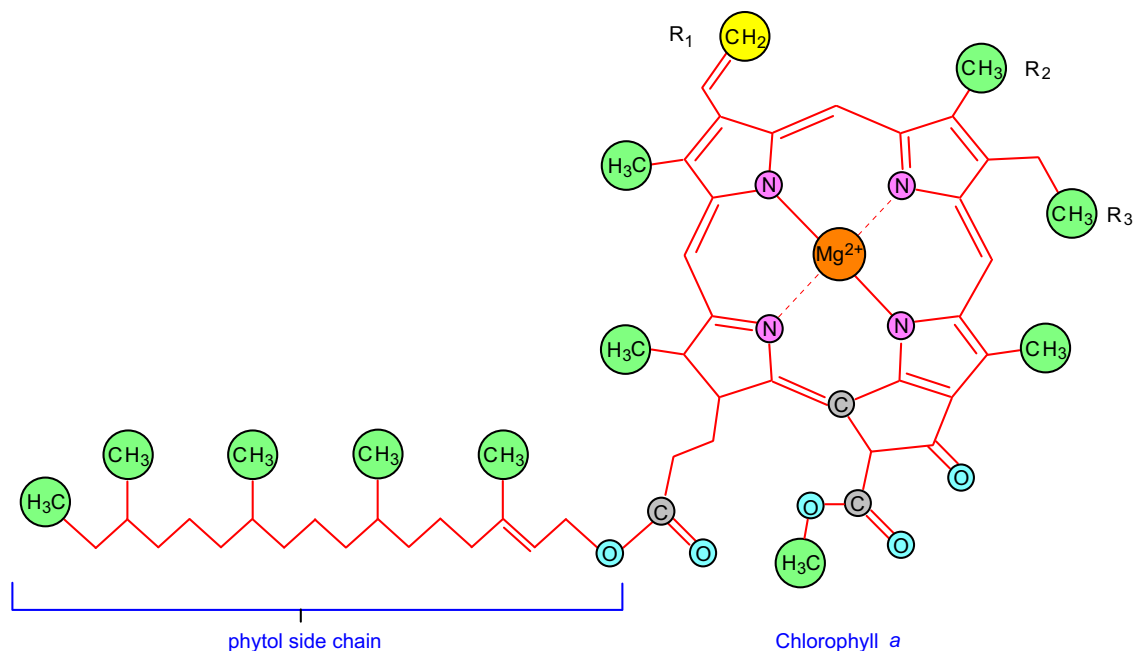


FIGURE 23.7 The structure of chlorophyll. Adapted from House (2013).

the ultraviolet region. Although not strictly aromatic conjugated ring systems, the phosphazines discussed in Chapter 14 also exhibit absorption in the ultraviolet region.

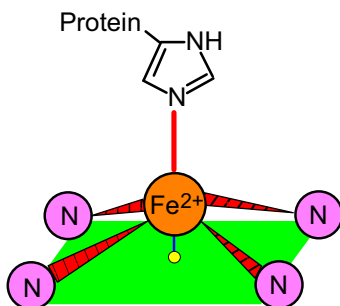
## 23.3 OXYGEN TRANSPORT

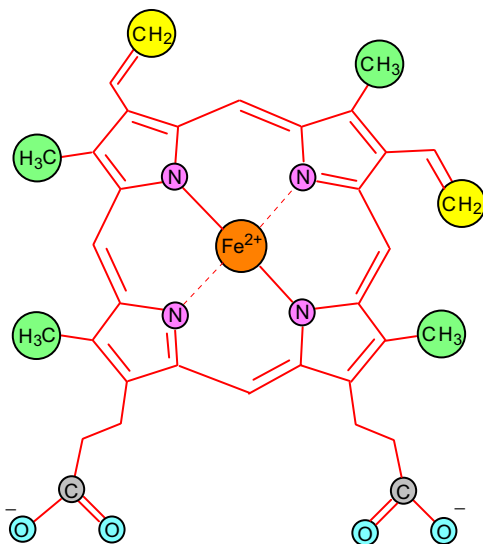
### 23.3.1 Hemoglobin

In the vast majority of species of fauna, oxygen transport is a vital part of the process of respiration. Oxygen transport is accomplished by iron in red blood cells. The versatility of iron to function in several biological processes is partially the result of ability of  $\text{Fe}^{2+}$  and  $\text{Fe}^{3+}$  to bond in different ways. For example, the ions have  $d^6$  and  $d^5$  configurations, respectively. Therefore, complexes containing the ions may be either high or low spin depending on the strength of the field generated by the ligands. There is also the possibility that when the metal ions are surrounded by four ligands that a tetrahedral bonding arrangement may exist. It is this ability of  $\text{Fe}^{2+}$  and  $\text{Fe}^{3+}$  to change bonding modes that is involved as oxygen molecules attach to iron in the heme structure.

In the heme environment, the iron is normally present as  $\text{Fe}^{2+}$  that is attached to the nitrogen atoms in the heme structure, which is shown in Figure 23.8. Although not shown in the figure, an imidazole ring is attached to the iron, which is in the high spin state with four unpaired electrons (see Chapter 20) and has a radius of 78 pm.

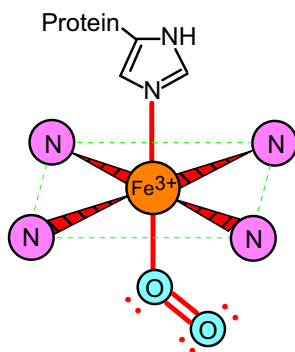
Although in the structure shown in Figure 23.8 the iron appears to lie in the same plane as the four nitrogen atoms, that is not the case when there is no  $\text{O}_2$  molecule attached (deoxyhemoglobin). In that case, the environment around the  $\text{Fe}^{2+}$  is as shown below with the  $\text{Fe}^{2+}$  residing about 60 pm (0.6 Å) above the plane of the four nitrogen atoms and the imidazole ring attached to the iron on the side above the plane of the four nitrogen atoms.





**FIGURE 23.8** The structure of heme with the imidazole group attached to iron not shown. Also, the protein to which the heme group is attached is omitted.

In this structure, the distance from the nitrogen atoms to the center of the square plane that they form is approximately 207 pm whereas the distances between the iron and the nitrogen atoms is 218 pm as a result of the iron not being located in the plane. When an oxygen molecule becomes attached, the iron, which will be shown as  $\text{Fe}^{3+}$ , lies in the same plane as the four nitrogen atoms. The groups bound to iron form an irregular octahedron.



Although electron density is transferred from the iron to the oxygen molecule, the exact nature of these species is unclear. The iron also changes to the low spin state in which the effective ionic radius is approximately 61 pm. Several values have been reported for the Fe–O distance in heme with the range being approximately 176–182 pm. However, in  $[\text{Fe}(\text{H}_2\text{O})_6]^{2+}$  there is a single Fe–O bond having a length of 209.5 pm.

There is some uncertainty regarding the actual oxidation state of the iron and the oxygen molecule in the oxygenated heme structure. If an electron is transferred from  $\text{Fe}^{2+}$  to the  $\text{O}_2$  molecule, the iron would exist as  $\text{Fe}^{3+}$ . The result would be the  $\text{O}_2^-$  (superoxide) ion. Examination of the molecular orbital diagram for  $\text{O}_2^-$  shows that the addition of an electron to a  $\pi^*$  orbital would result in a bond order of 1.5, intermediate between that of the  $\text{O}_2$  molecule (bond order = 2, bond length = 121 pm) and the peroxide ion,  $\text{O}_2^{2-}$  (bond order = 1, bond length 149 pm). As expected, bond length for the  $\text{O}_2^-$  ion is 126–133 pm, intermediate between the values for  $\text{O}_2$  and  $\text{O}_2^{2-}$ . This results in the O–O bond becoming weaker, which naturally results in it allowing bond cleavage with a lower energy requirement.

It should not be surprising that substances that bind more strongly than oxygen to the iron in heme would prevent the  $\text{O}_2$  molecules from binding. In that way, they destroy the oxygen carrying capability of heme. Carbon monoxide is one such gas that functions in that way and is thereby toxic, often with fatal results. Treatment of patients who have inhaled carbon monoxide involves placing them in an oxygen-rich environment to help displace CO molecules by Le Chatelier's principle.



### 23.3.2 Metal Ions in Superoxide Dismutase

In the human body (and in other mammals), superoxides are produced as a by-product of oxygen metabolism. Superoxides can produce cell damage, and they are converted to either normal oxygen,  $O_2$ , or hydrogen peroxide,  $H_2O_2$ , by the enzyme *superoxide dismutase*. There are several forms of *superoxide dismutase*, and they vary according to the metal involved at the reactive site and the way that the protein is folded. In humans, two of the enzymes contain copper as the reactive site where the catalysis takes place and the third contains manganese as the reactive portion of the enzyme. The enzyme that contains copper also contains zinc, but the zinc is not directly involved in the catalytic process, but instead helps control the orientation of the molecule. The major features of the bonding are shown in Figure 23.9.

There are two redox reactions involved that are catalyzed by *superoxide dismutase*. When manganese is the metal, the reactions can be written as



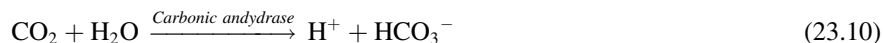
and the net result of the two reactions can be represented as



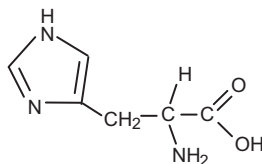
The corresponding redox reactions in which copper is part of the catalyst involves copper in the +1 and +2 oxidation states. Although hydrogen peroxide is also damaging to cells, it is converted to less damaging species by the enzyme *catalase*.

### 23.3.3 Zinc in Carbonic Anhydrase

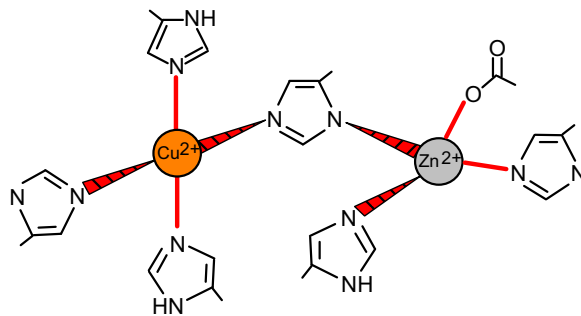
Zinc is an essential metal in several enzymes such as *aldolase*, *alcohol dehydrogenase*, and *alkaline phosphatase*. However, the enzyme *carbonic anhydrase* is responsible for the reaction



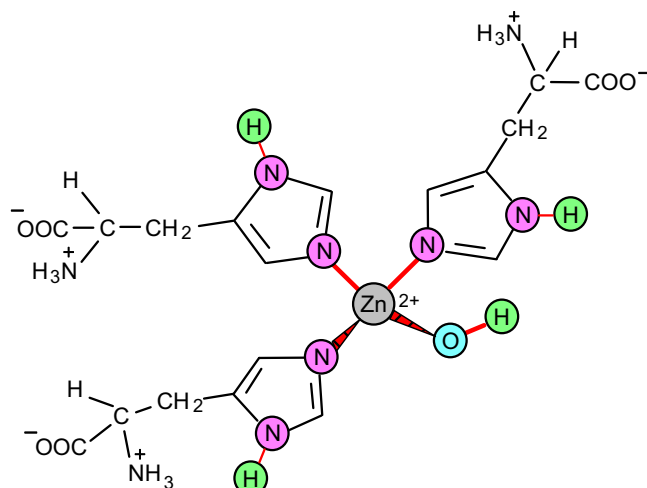
by which a relatively constant pH is maintained in the blood. The effective reaction site involves a zinc ion, and the enzyme is ineffective without it. In *carbonic anhydrase*, the zinc ion is surrounded by four histidine groups. Histidine is an amino acid that has the molecular structure



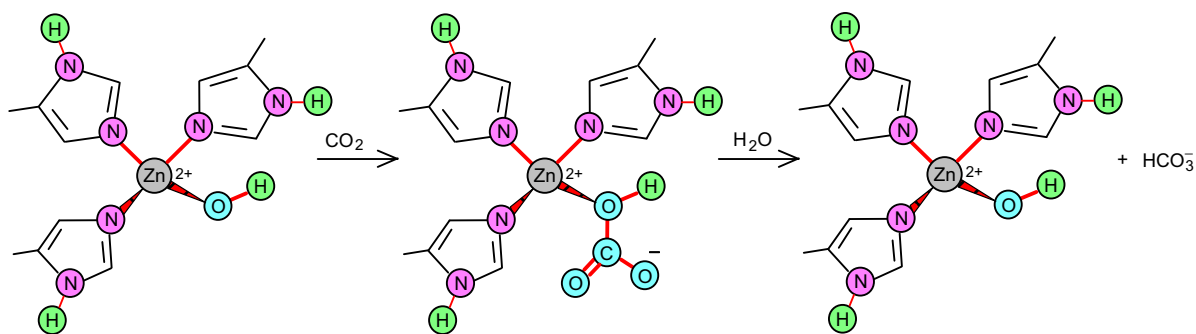
The bonding of zinc at the active site in *carbonic anhydrase* is shown in Figure 23.10.



**FIGURE 23.9** Bonding around the metal centers in *superoxide dismutase*. The distance between the metal ions is approximately 600 pm (6 Å). The bonding around the zinc is essentially tetrahedral whereas that around the copper is distorted.

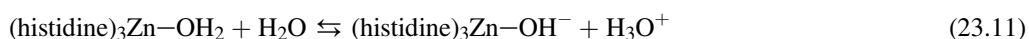


**FIGURE 23.10** The bonding around  $\text{Zn}^{2+}$  in the structure of *carbonic anhydrase*. This structure is attached to a protein, but only the active site is shown for clarity. The four bonds to zinc are essentially tetrahedral in orientation as would be expected for a  $d^{10}$  ion. Adapted from House (2013).

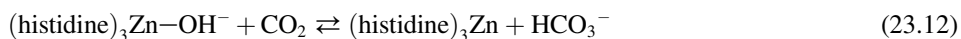


**FIGURE 23.11** A possible mechanism for the action of carbonic anhydrase. Adapted from House (2013).

As is the case for many reactions catalyzed by enzymes, the details of the mechanism of the reaction are not fully known. The first step apparently involves a reaction in which a water molecule coordinated to the zinc undergoes an ionization reaction that can be shown as



A general reaction of hydroxides is their reaction with  $\text{CO}_2$  to form bicarbonates and in this case the reaction can be shown as



The overall series of steps that accomplishes the transformation of  $\text{CO}_2$  into  $\text{HCO}_3^-$  is shown in [Figure 23.11](#). Substituting other metal ions for  $\text{Zn}^{2+}$  decreases activity of the catalyst and the presence of other potential ligands that bond to zinc (e.g.,  $\text{CN}^-$ ,  $\text{NO}_2^-$ , and  $\text{H}_2\text{S}$ ) causes the enzyme to lose its ability function as a catalyst.

The situations discussed in this chapter should be sufficient to show the breadth of those that depend on the presence of an inorganic compound or a structure containing a metal ion. A very large number of such processes are known to be important in biological systems, and they constitute a significant portion of the field known as bioinorganic chemistry. This is a burgeoning field, and the literature in the discipline is growing rapidly. Moreover, as some older applications of inorganic compounds fall into disuse, newer ones emerge. As a result, it can be surmised that this field will increase in importance in the future.

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## PROBLEMS

1. Although phosphoric acid is a weak acid, in aqueous solutions at normal pH the hydrogen ions are normally not present on the triphosphate chain. Explain this difference in acidity.
2. Write equations to show how  $\text{Fe}^{2+}$  or  $\text{Fe}^{3+}$  in soil would be converted into solid products that could not be utilized effectively by trees.
3. Interveinal chlorosis in oak trees, a disease caused by a deficiency of iron, often occurs when the trees grow in soil that has lime applied. Explain the chemistry involved.
4. The drug arsphenamine normally is obtained as the dihydrochloride. Would the protons be attached to the unshared pairs on arsenic or the amine groups? Explain your answer.
5. Iron in either the +2 or +3 state can form bonds in complexes in several ways. If  $\text{Fe}^{3+}$  forms a high spin octahedral complex, what would be the ligand field stabilization energy?
6. How would you expect the bond between two oxygen atoms to change when an  $\text{O}_2$  molecule becomes attached to iron in heme? What experimental techniques might be useful in determining the change in the bond?
7. Explain on the basis of electronic structure what type of complexes  $\text{Ag}^+$  would be most likely to form.
8. Why is catalysis of oxidation reactions by a transition metal ion more likely occur than for metal ions such as  $\text{Ca}^{2+}$  or  $\text{Mg}^{2+}$ ?
9. Suppose that  $\text{Al}^{3+}$  were to replace  $\text{Fe}^{2+}$  in a heme structure. Would you expect the reaction with oxygen to be the same? Explain your line of reasoning.
10. When an oxygen molecule is bound to iron in heme, the Fe–O bond length is approximately 180 pm in length. If the Fe–O bond length in  $[\text{Fe}(\text{H}_2\text{O})_6]^{2+}$  is 209.5 pm, what does that difference tell about the nature of the bond in the coordinated oxygen molecule?
11. Based on the supposed mode of action of salts of iron and manganese in treating itching caused by poison ivy (see Section 23.1.1), would you expect salts of calcium or magnesium to be equally effective? Explain your answer.

12. Calcium carbonate, magnesium hydroxide, sodium bicarbonate, and magnesium carbonate can be used to neutralize excess stomach acid (HCl). Write equations to show the reactions involved.
13. Write an equation to show how sodium fluoride reacts with hydroxyapatite in teeth to produce a substance that is more resistant to decay.
14. Draw the structure for the boric acid molecule. Explain the structural factors of this molecule that make it an acid that is so weak that it can be used as an eye wash.

## Appendix A

# Ground State Electron Configurations of Atoms

Atomic Number	Atom	Electron Configuration
1	H	$1s^1$
2	He	$1s^2$
3	Li	(He) $2s^1$
4	Be	(He) $2s^2$
5	B	(He) $2s^2 2p^1$
6	C	(He) $2s^2 2p^2$
7	N	(He) $2s^2 2p^3$
8	O	(He) $2s^2 2p^4$
9	F	(He) $2s^2 2p^5$
10	Ne	(He) $2s^2 2p^6$
11	Na	(Ne) $3s^1$
12	Mg	(Ne) $3s^2$
13	Al	(Ne) $3s^2 3p^1$
14	Si	(Ne) $3s^2 3p^2$
15	P	(Ne) $3s^2 3p^3$
16	S	(Ne) $3s^2 3p^4$
17	Cl	(Ne) $3s^2 3p^5$
18	Ar	(Ne) $3s^2 3p^6$
19	K	(Ar) $4s^1$
20	Ca	(Ar) $4s^2$
21	Sc	(Ar) $3d^1 4s^2$
22	Ti	(Ar) $3d^2 4s^2$
23	V	(Ar) $3d^3 4s^2$
24	Cr	(Ar) $3d^5 4s^1$
25	Mn	(Ar) $3d^5 4s^2$
26	Fe	(Ar) $3d^6 4s^2$
27	Co	(Ar) $3d^7 4s^2$
28	Ni	(Ar) $3d^8 4s^2$
29	Cu	(Ar) $3d^{10} 4s^1$
30	Zn	(Ar) $3d^{10} 4s^2$
31	Ga	(Ar) $3d^{10} 4s^2 4p^1$
32	Ge	(Ar) $3d^{10} 4s^2 4p^2$
33	As	(Ar) $3d^{10} 4s^2 4p^3$
34	Se	(Ar) $3d^{10} 4s^2 4p^4$
35	Br	(Ar) $3d^{10} 4s^2 4p^5$

(Continued)

—cont'd

Atomic Number	Atom	Electron Configuration
36	Kr	(Ar) $3d^{10} 4s^2 4p^6$
37	Rb	(Kr) $5s^1$
38	Sr	(Kr) $5s^2$
39	Y	(Kr) $4d^1 5s^2$
40	Zr	(Kr) $4d^2 5s^2$
41	Nb	(Kr) $4d^4 5s^1$
42	Mo	(Kr) $4d^5 5s^1$
43	Tc	(Kr) $4d^5 5s^2$
44	Ru	(Kr) $4d^7 5s^1$
45	Rh	(Kr) $4d^8 5s^1$
46	Pd	(Kr) $4d^{10}$
47	Ag	(Kr) $4d^{10} 5s^1$
48	Cd	(Kr) $4d^{10} 5s^2$
49	In	(Kr) $4d^{10} 5s^2 5p^1$
50	Sn	(Kr) $4d^{10} 5s^2 5p^2$
51	Sb	(Kr) $4d^{10} 5s^2 5p^3$
52	Te	(Kr) $4d^{10} 5s^2 5p^4$
53	I	(Kr) $4d^{10} 5s^2 5p^5$
54	Xe	(Kr) $4d^{10} 5s^2 5p^6$
55	Cs	(Xe) $6s^1$
56	Ba	(Xe) $6s^2$
57	La	(Xe) $5d^1 6s^2$
58	Ce	(Xe) $4f^2 6s^2$
59	Pr	(Xe) $4f^3 6s^2$
60	Nd	(Xe) $4f^4 6s^2$
61	Pm	(Xe) $4f^5 6s^2$
62	Sm	(Xe) $4f^6 6s^2$
63	Eu	(Xe) $4f^7 6s^2$
64	Gd	(Xe) $4f^7 5d^1 6s^2$
65	Tb	(Xe) $4f^9 6s^2$
66	Dy	(Xe) $4f^{10} 6s^2$
67	Ho	(Xe) $4f^{11} 6s^2$
68	Er	(Xe) $4f^{12} 6s^2$
69	Tm	(Xe) $4f^{13} 6s^2$
70	Yb	(Xe) $4f^{14} 6s^2$
71	Lu	(Xe) $4f^{14} 5d^1 6s^2$
72	Hf	(Xe) $4f^{14} 5d^2 6s^2$
73	Ta	(Xe) $4f^{14} 5d^3 6s^2$
74	W	(Xe) $4f^{14} 5d^4 6s^2$
75	Re	(Xe) $4f^{14} 5d^5 6s^2$
76	Os	(Xe) $4f^{14} 5d^6 6s^2$
77	Ir	(Xe) $4f^{14} 5d^7 6s^2$
78	Pt	(Xe) $4f^{14} 5d^9 6s^1$
79	Au	(Xe) $4f^{14} 5d^{10} 6s^1$

—cont'd

Atomic Number	Atom	Electron Configuration
80	Hg	(Xe) $4f^{14} 5d^{10} 6s^2$
81	Tl	(Xe) $4f^{14} 5d^{10} 6s^2 6p^1$
82	Pb	(Xe) $4f^{14} 5d^{10} 6s^2 6p^2$
83	Bi	(Xe) $4f^{14} 5d^{10} 6s^2 6p^3$
84	Po	(Xe) $4f^{14} 5d^{10} 6s^2 6p^4$
85	At	(Xe) $4f^{14} 5d^{10} 6s^2 6p^5$
86	Rn	(Xe) $4f^{14} 5d^{10} 6s^2 6p^6$
87	Fr	(Rn) $7s^1$
88	Ra	(Rn) $7s^2$
89	Ac	(Rn) $6d^1 7s^2$
90	Th	(Rn) $6d^2 7s^2$
91	Pa	(Rn) $5f^2 6d^1 7s^2$
92	U	(Rn) $5f^3 6d^1 7s^2$
93	Np	(Rn) $5f^5 7s^2$
94	Pu	(Rn) $5f^6 7s^2$
95	Am	(Rn) $5f^7 7s^2$
96	Cm	(Rn) $5f^7 6d^1 7s^2$
97	Bk	(Rn) $5f^8 6d^1 7s^2$
98	Cf	(Rn) $5f^{10} 7s^2$
99	Es	(Rn) $5f^{11} 7s^2$
100	Fm	(Rn) $5f^{12} 7s^2$
101	Md	(Rn) $5f^{13} 7s^2$
102	No	(Rn) $5f^{14} 7s^2$
103	Lr	(Rn) $5f^{14} 6d^1 7s^2$
104	Rf	(Rn) $5f^{14} 6d^2 7s^2$
105	Ha	(Rn) $5f^{14} 6d^3 7s^2$
106	Sg	(Rn) $5f^{14} 6d^4 7s^2$
107	Ns	(Rn) $5f^{14} 6d^5 7s^2$
108	Hs	(Rn) $5f^{14} 6d^6 7s^2$
109	Mt	(Rn) $5f^{14} 6d^7 7s^2$
110	Ds	(Rn) $5f^{14} 6d^8 7s^2$
111	Rg	(Rn) $5f^{14} 6d^9 7s^2$

## Appendix B

# Ionization Energies

Element	First Ionization Potential (kJ mol <sup>-1</sup> )	Second Ionization Potential (kJ mol <sup>-1</sup> )	Third Ionization Potential (kJ mol <sup>-1</sup> )
Hydrogen	1312.0	–	–
Helium	2372.3	5250.4	–
Lithium	513.3	7298.0	11,814.8
Beryllium	899.4	1757.1	14,848
Boron	800.6	2427	3660
Carbon	1086.2	2352	4620
Nitrogen	1402.3	2856.1	4578.0
Oxygen	1313.9	3388.2	5300.3
Fluorine	1681	3374	6050
Neon	2080.6	3952.2	6122
Sodium	495.8	4562.4	6912
Magnesium	737.7	1450.7	7732.6
Aluminum	577.4	1816.6	2744.6
Silicon	786.5	1577.1	3231.4
Phosphorus	1011.7	1903.2	2912
Sulfur	999.6	2251	3361
Chlorine	1251.1	2297	3826
Argon	1520.4	2665.2	3928
Potassium	418.8	3051.4	4411
Calcium	589.7	1145	4910
Scandium	631	1235	2389
Titanium	658	1310	2652
Vanadium	650	1414	2828
Chromium	652.7	1592	2987
Manganese	717.4	1509.0	3248.4
Iron	759.3	1561	2957
Cobalt	760.0	1646	3232
Nickel	736.7	1753.0	3393
Copper	745.4	1958	3554
Zinc	906.4	1733.3	3832.6
Gallium	578.8	1979	2963
Germanium	762.1	1537	3302
Arsenic	947.0	1798	2735
Selenium	940.9	2044	2974
Bromine	1139.9	2104	3500
Krypton	1350.7	2350	3565

(Continued)



—cont'd

Element	First Ionization Potential (kJ mol <sup>-1</sup> )	Second Ionization Potential (kJ mol <sup>-1</sup> )	Third Ionization Potential (kJ mol <sup>-1</sup> )
Rubidium	403.0	2632	3900
Strontium	549.5	1064.2	4210
Yttrium	616	1181	1980
Zirconium	660	1257	2218
Niobium	664	1382	2416
Molybdenum	685.0	1558	2621
Technetium	702	1472	2850
Ruthenium	711	1617	2747
Rhodium	720	1744	2997
Palladium	805	1875	3177
Silver	731.0	2073	3361
Cadmium	867.6	1631	3616
Indium	558.3	1820.6	2704
Tin	708.6	1411.8	2943.0
Antimony	833.7	1794	2443
Tellurium	869.2	1795	2698
Iodine	1008.4	1845.9	3200
Xenon	1170.4	2046	3097
Cesium	375.7	2420	—
Barium	502.8	965.1	—
Lanthanum	538.1	1067	—
Cerium	527.4	1047	1949
Praseodymium	523.1	1018	2086
Neodymium	529.6	1035	2130
Promethium	535.9	1052	2150
Samarium	543.3	1068	2260
Europium	546.7	1085	2404
Gadolinium	592.5	1167	1990
Terbium	564.6	1112	2114
Dysprosium	571.9	1126	2200
Holmium	580.7	1139	2204
Erbium	588.7	1151	2194
Thulium	596.7	1163	2285
Ytterbium	603.4	1176	2415
Lutetium	523.5	1340	2022
Hafnium	642	1440	2250
Tantalum	761	(1500)	—
Tungsten	770	(1700)	—
Rhenium	760	1260	2510
Osmium	840	(1600)	—
Iridium	880	(1680)	—
Platinum	870	1791	—
Gold	890.1	1980	—

—cont'd

Element	First Ionization Potential (kJ mol <sup>-1</sup> )	Second Ionization Potential (kJ mol <sup>-1</sup> )	Third Ionization Potential (kJ mol <sup>-1</sup> )
Mercury	1007.0	1809.7	3300
Thallium	589.3	1971.0	2878
Lead	715.5	1450.4	3081
Bismuth	703.2	1610	2466
Polonium	812	(1800)	—
Astatine	930	1600	—
Radon	1037	—	—
Francium	400	(2100)	—
Radium	509.3	979.0	—
Actinium	499	1170	—
Thorium	587	1110	—
Protactinium	568	—	—
Uranium	584	1420	—
Neptunium	597	—	—
Plutonium	585	—	—
Americium	578.2	—	—
Curium	581	—	—
Berkelium	601	—	—
Californium	608	—	—
Einsteinium	619	—	—
Fermium	627	—	—
Mendelevium	635	—	—
Nobelium	642	—	—

Numbers in parentheses are approximate values.

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