MODULAR SYSTEM

CHEMICAL BONDS

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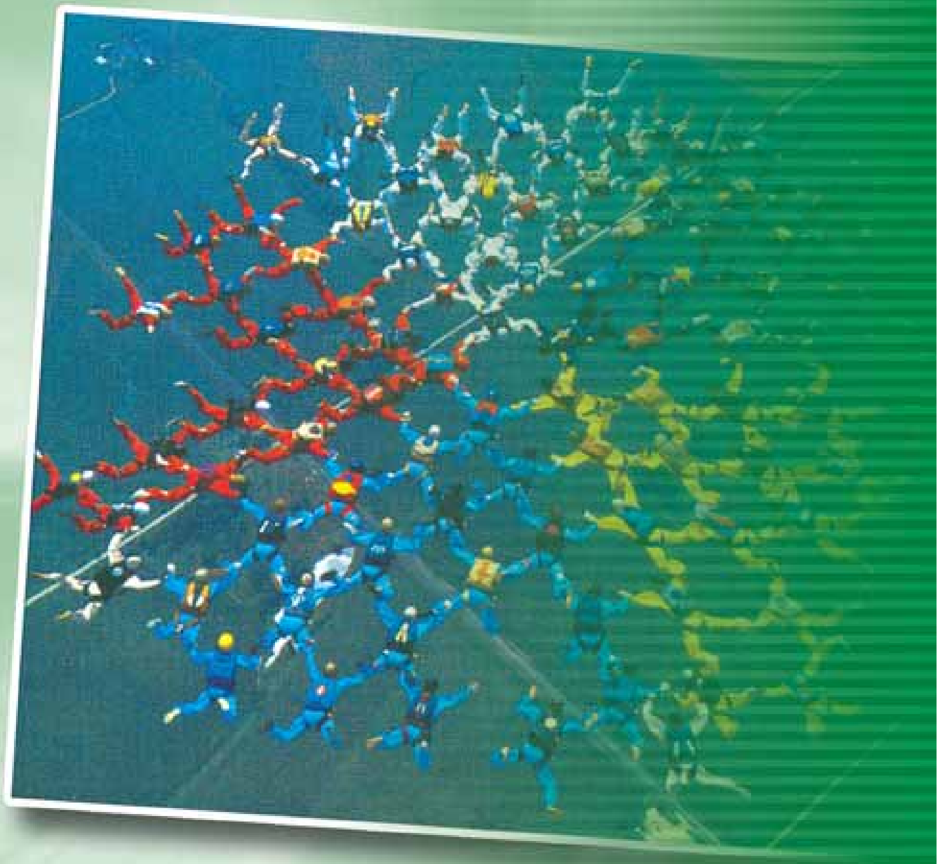

PREFACE

Chemistry is an interesting and fundamental branch of science because it gives us the chance to explain the secrets of nature. What is water? What do we use in our cars as a fuel? What is aspirin? What are perfumes made of? Many of these kind of questions and their answers are all part of the world of chemistry. There is no industry that does not depend upon chemical substances: petroleum, pharmaceuticals, garment, aircraft, steel, electronics, agricultural, etc. This book helps everyone to understand nature. However, one does not need to be a chemist or scientist to understand the simplicity within the complexity around us.

The aim was to write a modern, up-to-date book where students and teachers can get concise information about the structure of substances. Sometimes reactions are given in detailed form, but, in all, excessive detail has been omitted.

The book is designed to introduce basic knowledge about chemical bonds. Chemists work everyday to produce new compounds to make our lives easier with the help of this basic knowledge. In the design, emphasis has been placed upon making the book student friendly. Throughout the books, colorful tables, important reactions, funny cartoons, interesting extras and reading passages are used to help explain ideas.

The authors would like to thank Orhan Keskin, Ali Çavdar and Ramazan Şahin for their support and encouragement throughout the development of this book.

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We are particularly grateful to our spouses and children for their patience during the writing of the book.

The Authors

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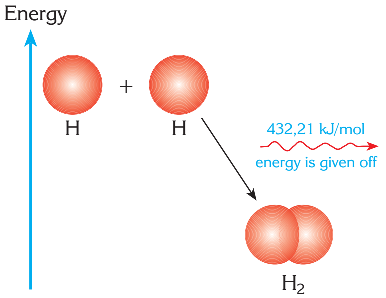
CHEMICAL BONDS

INTRODUCTION

Atoms and ions are generally found connected together in groups. How are these groups held together and what is responsible for bringing about the distinctive properties of a substance?

What are the reasons that cause iron to be solid, water to be liquid and hydrogen to be a gas at room temperature? Why is diamond hard while wax is soft? Why do some solids melt at low temperatures while others melt at high temperatures?

For example, carbon and silicon are found within the same group in the periodic table. Considering the trends in a group, we would expect the oxides of these two elements, CO_2 and SiO_2, to display similar properties. However, SiO_2 is a solid with a quartz structure while CO_2 is a gas that has great importance in the life cycle. What can be the reason for these two compounds being so different?



The energy of a molecule is usually less than the energy of the atoms that form the molecule. So when atoms form a molecule, they give off energy and become more stable.

Substances	Melting Point (°C)
Oxygen (O_2)	–218
Carbon dioxide (CO_2)	–56
Ice (H_2O)	0
Sand (SiO_2)	1640
Iron (Fe)	1535
Tungsten (W)	3410

Table 1: *Melting points of some substances*

We are going to study the answers to these questions in this module.

We can represent groups of atoms or ions by models. For example, to hold two ping-pong balls together they must be stuck or connected by a rod. There likewise must be a connection between the sodium and chloride ions in table salt or between the hydrogen and oxygen atoms in water. This force of attraction that holds atoms or ions together is called a ***chemical bond***.

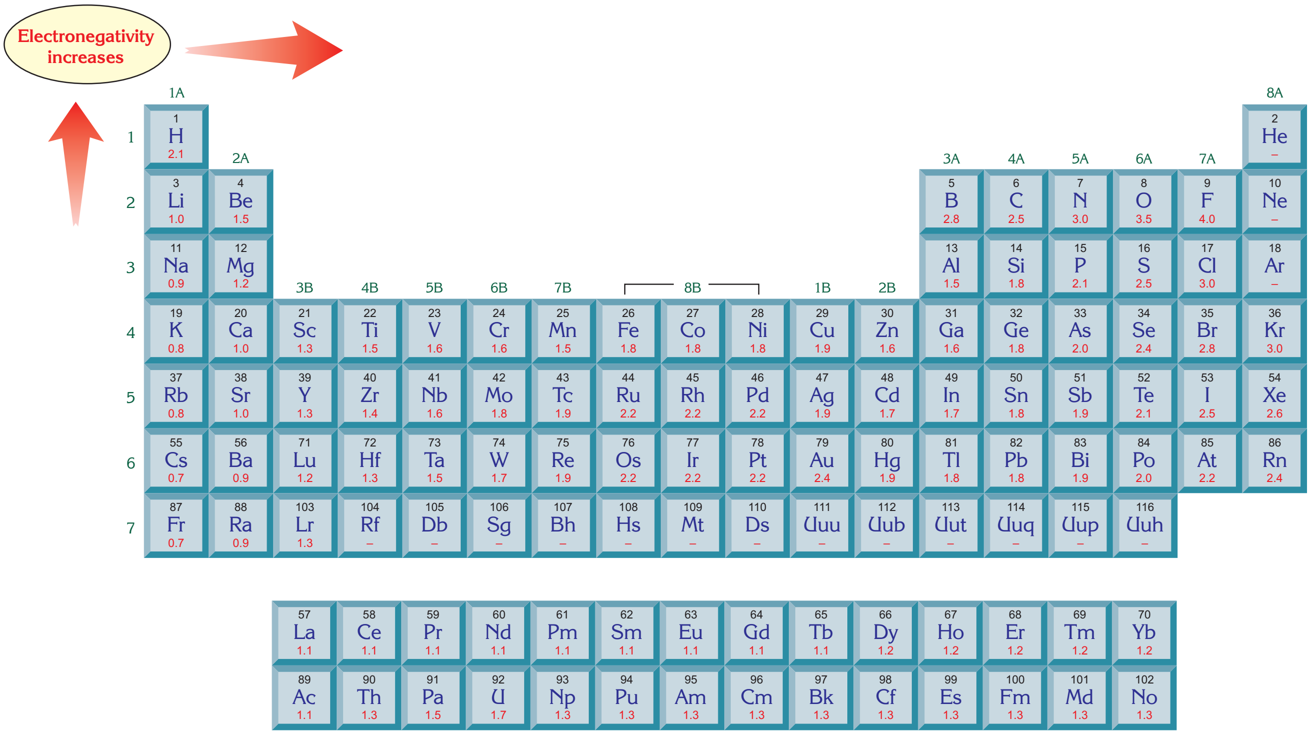
A knowledge of chemical bonds is important to help us to understand chemical reactions. In a chemical reaction, bonds are broken and new bonds are formed. During this process the total energy of the substances changes. For example, the energy of a molecule is generally less than that of the individual atoms that make up the molecule.

During the process of forming a chemical bond, energy is given out, and this energy is equal to that required to break the same chemical bond. To gain a better understanding of chemical bonds we need to study electronegativity, a property that plays an important role in bond formation.

1. ELECTRONEGATIVITY

Electronegativity is a term which was first proposed in 1934 by the American physicist R.S. Mulliken. It is especially useful in explaining the type of the bond occurring between atoms.

Electronegativity is the tendency of an atom to attract the bonding electrons within a compound to itself. It depends upon the ***nuclear charge*** (proton number) and the ***atomic radius*** of the atom. It is these factors that control the ionization energy of the atom which in turn is related to the ability of an atom to attract electrons.



	1A	2A	3B	4B	5B	6B	7B	8B			1B	2B	3A	4A	5A	6A	7A	8A
1	1 H 2.1																	2 He –
2	3 Li 1.0	4 Be 1.5											5 B 2.8	6 C 2.5	7 N 3.0	8 O 3.5	9 F 4.0	10 Ne –
3	11 Na 0.9	12 Mg 1.2											13 Al 1.5	14 Si 1.8	15 P 2.1	16 S 2.5	17 Cl 3.0	18 Ar –
4	19 K 0.8	20 Ca 1.0	21 Sc 1.3	22 Ti 1.5	23 V 1.6	24 Cr 1.6	25 Mn 1.5	26 Fe 1.8	27 Co 1.8	28 Ni 1.8	29 Cu 1.9	30 Zn 1.6	31 Ga 1.6	32 Ge 1.8	33 As 2.0	34 Se 2.4	35 Br 2.8	36 Kr 3.0
5	37 Rb 0.8	38 Sr 1.0	39 Y 1.3	40 Zr 1.4	41 Nb 1.6	42 Mo 1.8	43 Tc 1.9	44 Ru 2.2	45 Rh 2.2	46 Pd 2.2	47 Ag 1.9	48 Cd 1.7	49 In 1.7	50 Sn 1.8	51 Sb 1.9	52 Te 2.1	53 I 2.5	54 Xe 2.6
6	55 Cs 0.7	56 Ba 0.9	71 Lu 1.2	72 Hf 1.3	73 Ta 1.5	74 W 1.7	75 Re 1.9	76 Os 2.2	77 Ir 2.2	78 Pt 2.2	79 Au 2.4	80 Hg 1.9	81 Tl 1.8	82 Pb 1.8	83 Bi 1.9	84 Po 2.0	85 At 2.2	86 Rn 2.4
7	87 Fr 0.7	88 Ra 0.9	103 Lr 1.3	104 Rf –	105 Db –	106 Sg –	107 Bh –	108 Hs –	109 Mt –	110 Ds –	111 Uuu –	112 Uub –	113 Uut –	114 Uuq –	115 Uup –	116 Uuh –		

57 La 1.1	58 Ce 1.1	59 Pr 1.1	60 Nd 1.1	61 Pm 1.1	62 Sm 1.1	63 Eu 1.1	64 Gd 1.1	65 Tb 1.1	66 Dy 1.2	67 Ho 1.2	68 Er 1.2	69 Tm 1.2	70 Yb 1.2
89 Ac 1.1	90 Th 1.3	91 Pa 1.5	92 U 1.7	93 Np 1.3	94 Pu 1.3	95 Am 1.3	96 Cm 1.3	97 Bk 1.3	98 Cf 1.3	99 Es 1.3	100 Fm 1.3	101 Md 1.3	102 No 1.3

Figure 1: *Electronegativity values of the elements.*
In the periodic table, electronegativity increases from left to right and from bottom to top.

Electronegativity does not have a unit. The most widely used electronegativity scale today was derived by Linus Pauling. He used bond energy values in the preparation of this scale.

According to this scale the most active metal francium and the most active non-metal fluorine have electronegativity values of 0.7 and 4.0 respectively. In general, the electronegativity values of metals are smaller than 1.7 and those of non-metals greater than 1.8. Some noble gases do not have an electronegativity value as they don't form bonds with other elements. Electronegativity values of the elements are displayed in Figure 1.

As it is mentioned, electronegativity is dependent upon atomic radius. In the periodic table, as a period is crossed from left to right, atomic radius decreases, and hence the ability of an atom to attract valence electrons increases. However, as you descend a group, atomic radius increases and therefore ability of an atom to attract valence electrons decreases. So consequently, electronegativity decreases from top to bottom in a group and increases from left to right across a period.

EXTRA

Bond Polarity

The concept of bond polarity explains the behaviour of atoms how they share the bonding electrons between each other. Electronegativity difference of the atoms expresses bond polarity in a molecule.

For example, the electronegativity differences between the atoms are:

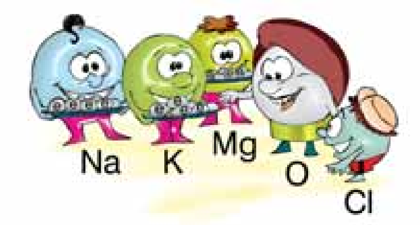
Li(1.0) – Cl(3.0) = 2

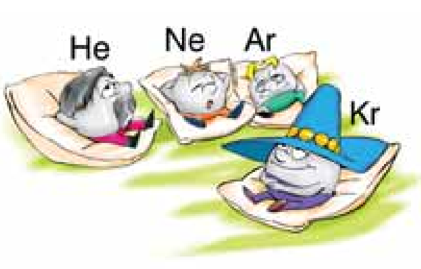
Mg(1.2) – Cl(3.0) = 1.8

H(2.1) – Cl(3.0) = 0.9

As a result;

The highest electronegativity difference is between lithium and chlorine and the lowest difference is between hydrogen and chlorine. Therefore, the bond between Li and Cl (Li–Cl) is the most polar.





2. CHEMICAL BONDS AND THEIR FORMATION

When atoms get closer to each other, they may become held together by forces of attraction called chemical bonds. To explain why this happens, we need to understand more about the electron configurations of atoms.

The noble gases (He, Ne, Ar, Kr, Xe and Rn) which form group 8A in the periodic table are the most stable elements. They all have the ns^2 np^6 electron configuration (except He which has the $1s^2$ configuration). The other elements in the periodic table have a tendency to gain the electron configuration of a noble gas and hence become stable. For this reason, atoms want to complete their last shell and gain the ns^2 np^6 configuration.

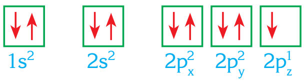
The tendency of atoms to make the number of their valence electrons eight, like the nobel gases, is known as the ***octet rule***. There are two ways for the elements to gain their octet and obtain a noble gas electron configuration.

1. Electron transfer
2. Electron sharing

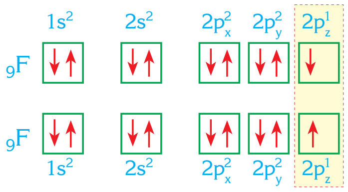
Chemical bonds are classified into two groups; transfer of electrons creates an ionic bond while the sharing of electrons leads to a covalent bond. Before studying chemical bonds we need to become familiar with their representation. Chemical bonds may be represented in several ways. We are going to study orbital representation, electron dot representation and line representation. Let's examine these three types using the example of the fluorine molecule, F_2.

Orbital Representation

The electron configuration of fluorine is $1s^22s^22p^5$ and its orbital representation is;



Two fluorine atoms join together to increase their number of valence electrons to eight. When their half - filled $2p_z$ orbitals overlap, a bond is formed. As a result, each fluorine atom completes its octet and together they form the stable fluorine molecule.



Electron Dot Representation (Lewis Symbol)

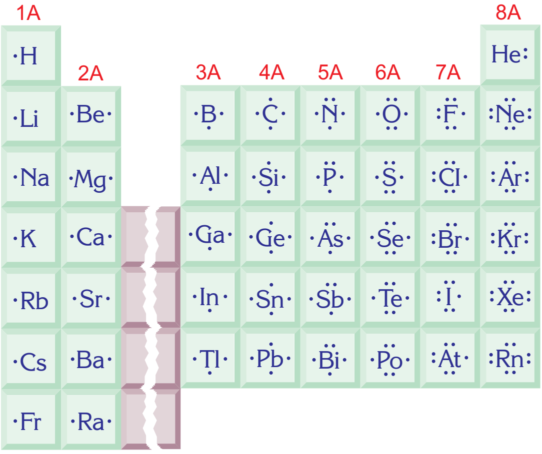
This representation is also known as the Lewis symbol representation. In this representation valence electrons are shown as dots around the symbol of the element.

When we look at the electron configuration of the fluorine atom, we see that it has seven valence electrons. Therefore the electron dot representation of fluorine atom is $\cdot\underset{\cdot\cdot}{\ddot{F}}:$ though it can also be represented by $:\underset{\cdot}{\ddot{F}}:$, $:\underset{\cdot\cdot}{\ddot{F}}\cdot$ or $:\underset{\cdot\cdot}{\dot{F}}:$

When two fluorine atoms combine with each other a F_2 molecule forms.

$$:\underset{\cdot\cdot}{\ddot{F}}\cdot + \cdot\underset{\cdot\cdot}{\ddot{F}}: \longrightarrow :\underset{\cdot\cdot}{\ddot{F}}:\underset{\cdot\cdot}{\ddot{F}}:$$

The electron pair ":" between two fluorine atoms ($:\underset{\cdot\cdot}{\ddot{F}}\textcircled{:}\underset{\cdot\cdot}{\ddot{F}}:$) represents the bond and other electron pairs represent unbonded electron pairs.

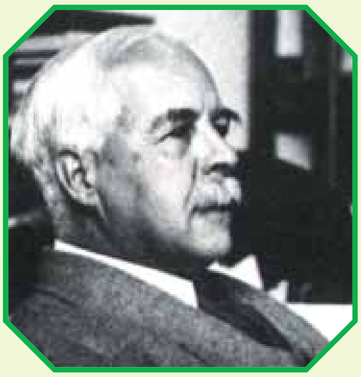


Lewis symbols of group A elements.

Line Representation

Bond structure can also be represented by lines. Each electron pair is shown by a line. In other words two electrons ":" are shown by a line "–". So the line representation of the fluorine molecule is $|\overline{\underline{F}}-\overline{\underline{F}}|$. The line between the two fluorine atoms represents the bond. Sometimes both the Lewis symbol and line representation can be used in the same molecule. For example, the F_2 molecule can also be represented as $:\underset{\cdot\cdot}{\ddot{F}}-\underset{\cdot\cdot}{\ddot{F}}:$

THE PIONEERS



Gilbert Newton Lewis
(1875 - 1946)

Lewis was an American scientist born in 1875 in Massachusetts, USA. He started his academic career in 1912 and proposed the theory of electron sharing in 1916 which as we have seen is of great importance to chemists. Because of this theory, "electron dot representation" is also named "Lewis dot structure".

Besides chemical bonds, Lewis also studied thermodynamics, isotopes and light. He expanded his theories of chemical bonding and also proposed the Lewis acid-base theory.

READING

HOW TO WRITE LEWIS STRUCTURES OF MOLECULES

Have you ever wondered how to draw the structures of compounds? For example, compounds such as CCl_4, PBr_3 or ions such as SO_4^{2-}. To draw the structural formulae we will use the Lewis (electron dot) notation.

To learn how to construct Lewis structures for molecules let's examine the following rules by applying them to the CCl_4 molecule;

1. Determine the total number of valence electrons in the molecule.

The carbon atom is in group 4A of the periodic table, so it has 4 valence electrons and chlorine is in 7A group, so it has 7 valence electrons.

1 C atom : $1 \cdot 4 = 4$ valence electrons

4 Cl atoms : $4 \cdot 7 = 28$ electrons

Total number of valence electrons : $4 + 28 = 32$

2. Determine the total number of electrons needed to complete the octets (the number of valence electrons should be 8, though for hydrogen it is 2) for the atoms.

1 C atom + 4 Cl atoms = 5 atoms

$5 \cdot 8 = 40$ electrons

3. Subtract the number of electrons you obtained in step two from that of step 1.

This difference gives us the number of electrons that are going to be used in bond formation.

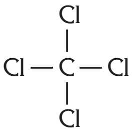
$40 - 32 = 8$ electrons are going to be used in bond formation.

4. To form a bond 2 electrons are needed.

So from step three, $8 / 2 = 4$ bonds are going to be formed.

5. Identify the central atom. This is most often the atom present with the lowest number. Write the skeleton structure and then join the atoms by single covalent bonds.

In the CCl_4 molecule, carbon is the centralatom, because it has the lowest number (1C, 4Cl).

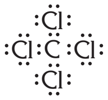


$$\begin{array}{c} \mathrm{Cl} \\ | \\ \mathrm{Cl}-\mathrm{C}-\mathrm{Cl} \\ | \\ \mathrm{Cl} \end{array}$$

6. For each single bond formed, subtract two electrons from the total number of valence electrons. Thus the electrons needed to complete the octet of each atom in the molecule are found.

In the CCl_4 molecule the number of total valence electrons is 32. 8 electrons are used in bond formation. So $32 - 8 = 24$ electrons remain.

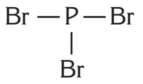
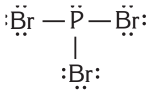
Those remaining electrons are distributed around the chlorine atoms in a manner that would give each chlorine atom 8 electrons. So 6 more electrons are needed for each chlorine atom to complete its octet.



Thus each chlorine atom has 8 electrons and carbon also has 8 electrons (4 bonds $\cdot$ 2 = 8 electrons)

Example 1:

Write the Lewis structure of the PBr_3 molecule.

1. *P atom is in group 5A and Br atom is in group 7A.*
   1 P atom contains 5 valence electrons
   3 Br atoms contain $3 \cdot 7 = 21$ valence electrons.
   Total number of valence electrons = 26
2. *1 P atom + 3 Br atoms = 4 atoms*
   $4 \cdot 8 = 32$ electrons.
3. *$32 - 26 = 6$ electrons are used in bond formation.*
4. *$6 : 2 = 3$ bonds are formed.*
5. 
   $Br-\underset{\underset{Br}{|}}{P}-Br$
6. *$26 - 6 = 20$ electrons should be distributed around the atoms so that each would have 8 electrons.*
7. 
   $:\underset{..}{\ddot{Br}}-\ddot{P}-\underset{..}{\ddot{Br}}:$ with $\underset{:\underset{..}{Br}:}{|}$ below P

Writing Lewis Structure

1. *Determine the total number of valence electrons*

⇩

2. *Determine the total number of electrons needed to complete the octets*

⇩

3. *Subtract the number of electrons you obtained in step two from that of step one.*

⇩

4. *Divide the result of the step three by 2 to get the number of bonds.*

⇩

5. *Determine the central atom and then draw the bonds between atoms.*

⇩

6. *Subtract (number of bonds · 2) from the number of valence electrons to find nonbonding electrons.*

⇩

7. *Distribute the remaining electrons to atom to get octet.*

Example 1

For the HF molecule, show its;

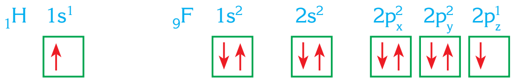
a. orbital representation,

b. electron dot representation,

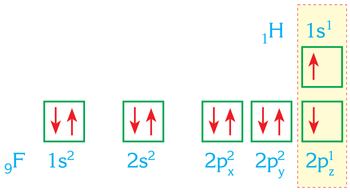
c. line representation ($_1H$, $_9F$)

Solution

a. First show the electron configurations and orbital representations of the hydrogen and fluorine atoms.

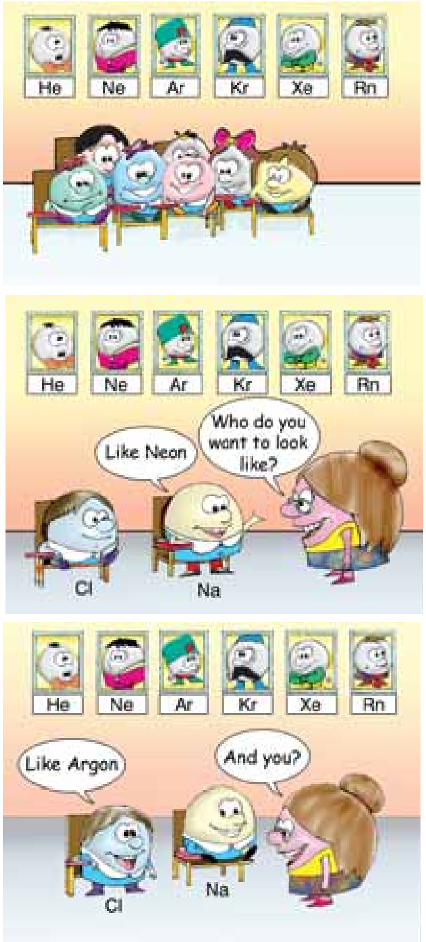


It is seen from their orbital structures that hydrogen and fluorine both need to share 1 electron to complete their outer shells. Therefore the orbital representation of HF molecule is;



b. The valence electrons of H and F are 1 and 7 respectively. So the Lewis symbols of H and F are $\mathrm{H}\cdot$ and $\cdot\ddot{\underset{\cdot\cdot}{\mathrm{F}}}\colon$. Therefore the electron dot representation of HF molecule is $\mathrm{H}\colon\ddot{\underset{\cdot\cdot}{\mathrm{F}}}\colon$

c. Since the electron dot structure of HF is $\mathrm{H}\colon\ddot{\underset{\cdot\cdot}{\mathrm{F}}}\colon$, the line representation is simply $\mathrm{H}-\overline{\underline{\mathrm{F}}}|$.



While forming compounds, all atoms tend to acquire noble gas electronic configuration.

2.1. IONIC BONDS

Ionic bonds are formed by the transfer of electrons from one atom to another. After the transfer of electrons, the atom that lost electrons becomes positively charged and the atom that gained electrons becomes negatively charged. The force of attraction that holds these atoms together is the electrostatic force between their opposite charges.

Ionic bonds are formed between atoms that have an electronegativity difference greater than about 1.9.

Let's consider the bond formation between sodium and chlorine, a metal and a nonmetal. The electronegativity values of sodium and chlorine are 0.9 and 3.0 respectively. This tells us that sodium has a low ionization energy and a tendency to give electrons while chlorine has a tendency to take electrons.

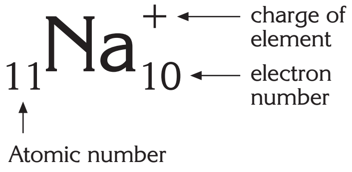
When those two atoms come together under suitable conditions, to complete their octets, sodium gives one electron to chlorine.

$_{11}\mathrm{Na}$: $1s^2\ 2s^2\ 2p^6\ 3s^1$

$_{17}\mathrm{Cl}$: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$

$${}_{11}\cdot\text{Na}_{11} \longrightarrow {}_{11}\text{Na}^{+}_{10} + e^{-} \quad \text{and} \quad {}_{17}\cdot\ddot{\underset{\cdot\cdot}{\text{Cl}}}:_{17} + e^{-} \longrightarrow {}_{17}:\ddot{\underset{\cdot\cdot}{\text{Cl}}}:^{-}_{18}$$

So the Na^{+} cation and Cl^{-} anion are formed.



Sodium loses its valence electron and its electron configuration becomes identical to that of neon: $1s^2\ 2s^2\ 2p^6$. Likewise, the valence shell of chlorine becomes completely filled and its electron configuration resembles that of argon. As a result, during the reaction

$$\text{Na}\cdot + \cdot\ddot{\underset{\cdot\cdot}{\text{Cl}}}: \longrightarrow \text{Na}^{+}[:\ddot{\underset{\cdot\cdot}{\text{Cl}}}:]^{-}$$

an ionic bond is formed between the sodium and chloride ions.

The Na^{+} and Cl^{-} ions can be considered as negatively and positively charged spheres that attract each other. Since positive (+) and negative (–) charges form an electric field in all directions, the electrostatic force of attraction (ionic bond) is not just in one direction. In the NaCl crystal, each Na^{+} ion is surrounded by six Cl^{-} ions and each Cl^{-} ion is surrounded by six Na^{+} ions (Figure 2). Because of this, the structure of NaCl is not a molecule but it is in the form of an ionic crystal in which many ions are found together.

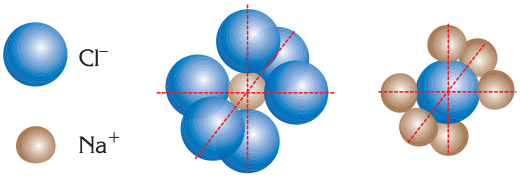
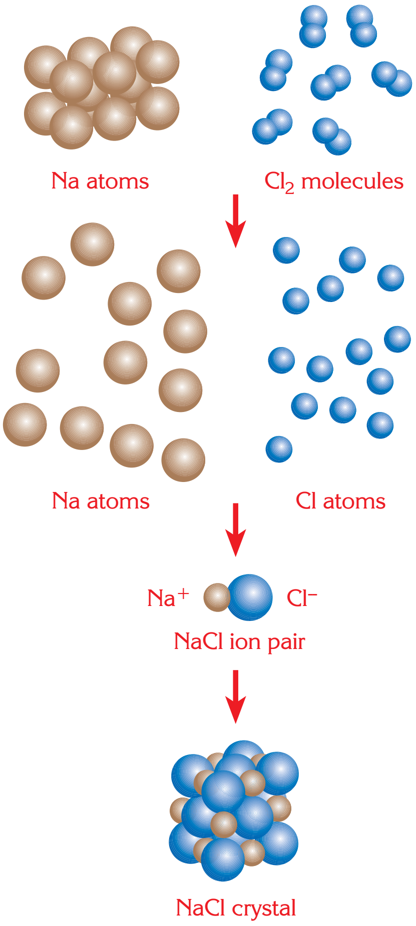


Figure 2: *In the sodium chloride crystal, each sodium ion is surrounded by six chloride ions and each chloride ion is surrounded by six sodium ions.*



Formation of the ionic bond between sodium and chloride ions in the NaCl crystal.

The degree of polarity of the bond is proportional to the electronegativity differences between the atoms. Because of this fact when the electronegativity difference between the atoms is large, as it is between most metals and non-metals, ionic bonding is the result.

Structures that contain ionic bonds are found in solid phase at room temperature.

A salt lake

Example 2

Compare the ionic character (polarity) of the bonds in NaBr and NaF.

The electronegativity values of the given elements are;

Na : 0.9, Br : 2.8, F : 4.0

Solution

The electronegativity difference between the Na atom and the Br atom is

2.8 – 0.9 = 1.9

The electronegativity difference between the Na atom and the F atom is

4 – 0.9 = 3.1

As the electronegativity difference in NaF is greater, the bond is more ionic than in NaBr.

As a result, we see that the electronegativity difference between Na and F is greater than that of Na and Br. Therefore the polarity of the bond in NaF is greater than that of NaBr.

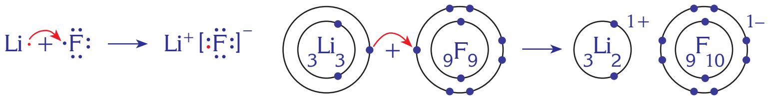
Example 3

Show the formation of ionic bonds between the following pairs.

a. ($_3$Li, $_9$F) b. ($_{20}$Ca, $_{35}$Br) c. ($_{13}$Al, $_8$O)

Solution

a. The electron dot structures of lithium and fluorine are Li· and ·F: so lithium has 1 valence electron and fluorine has 7 valence electrons. To achieve stability, lithium gives its single valance electron to fluorine and Li^+ and F^- ions are formed.



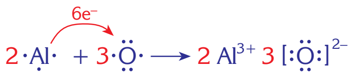
Therefore lithium fluoride, LiF, is formed.

b. The electron dot structures of the calcium atom and bromine atom are ·Ca· and · Br : respectively. It is seen that calcium has 2 valence electrons and bromine has 7 valence electrons. To achieve stability, calcium loses 2 electrons and forms the Ca^{2+} ion. For bromine to complete its octet it needs 1 electron. Therefore each calcium atom should bond with 2 bromine atoms.

As a result the compound calcium bromide with the formula $CaBr_2$ is formed.

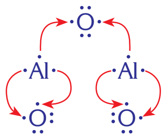
c. The Lewis dot structures of the aluminum and oxygen atoms are $\cdot\underset{\cdot}{\text{Al}}\cdot$ and $\cdot\overset{\cdot\cdot}{\underset{\cdot\cdot}{\text{O}}}\cdot$ respectively.

Aluminum has a tendency to lose its 3 outer electrons to form the Al^{3+} ion and oxygen has a tendency to gain 2 electrons to form O^{2-} ion. The aluminum atom has 3 valence electrons and oxygen atom needs 2 electrons to complete its octet so two aluminum atoms (in total 6 electrons are lost) form bonds with three oxygen atoms (in total 6 electrons are gained).



$$2\cdot\underset{\cdot}{\text{Al}}\cdot + 3\cdot\overset{\cdot\cdot}{\underset{\cdot\cdot}{\text{O}}}\cdot \longrightarrow 2\,Al^{3+}\;3\,[:\overset{\cdot\cdot}{\underset{\cdot\cdot}{\text{O}}}:]^{2-}$$

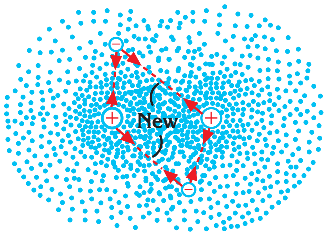
As a result, aluminum oxide with the formula Al_2O_3 is formed.



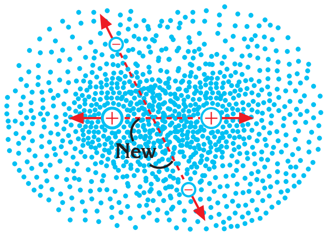
2.2. COVALENT BONDS

We know that the electronegativity difference between atoms must be greater than 1.9 to form an ionic bond. But if the electronegativity values of the atoms are similar, the tendency of the atoms to take or give electrons will also be similar. The transfer of electrons is not possible between such atoms, so the atoms must share electrons to gain a stable octet. The bond that is formed as a result of electron sharing is called a ***covalent bond***. Covalent bonds are generally formed between two nonmetals.

Let's examine the formation of a hydrogen molecule from two hydrogen atoms. Since the electron configuration of a hydrogen atom is $1s^1$, it must gain one more electron to reach the configuration of a noble gas (the $1s^2$ configuration of the He atom). So both hydrogen atoms which will form the hydrogen molecule need to take one more electron to be stable. Since there is no electronegativity difference, none of them can take an electron from the other. Instead the hydrogen atoms share their electrons and a covalent bond forms. The electron pair is attracted by the nuclei (protons) of both hydrogen atoms (Figure 3a). But counter to this the electrons and the two nuclei repel each other due to their similar charges (Figure 3b). Overall, the attractive and repulsive forces cancel out and in this state, the energy of the hydrogen molecule is less than the total energies of hydrogen atoms. In other words, the molecule is more stable than its constituent atoms.

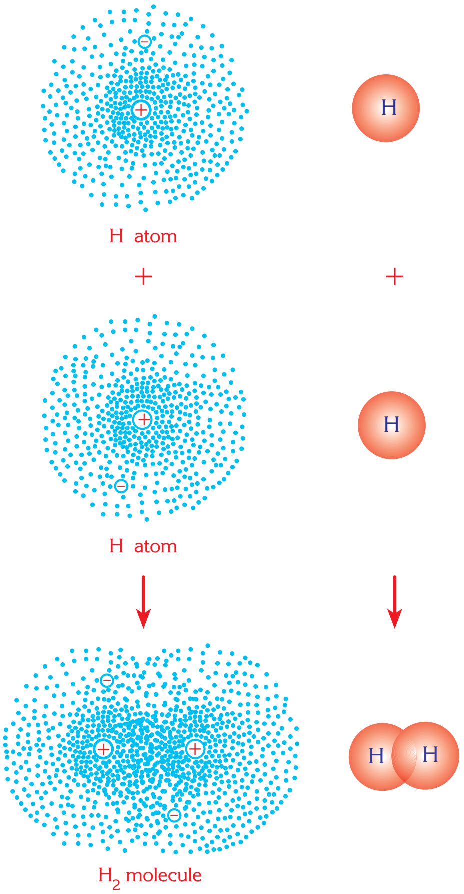


a. Attractive forces
(between proton and electron)



b. Repulsive forces
(between proton - proton and
electron - electron)

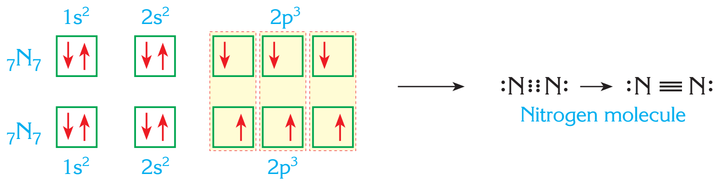
Figure 3: *Formation of the covalent bond between hydrogen atoms. When two atoms are far from each other, the electrons of the atom are attracted only by the nucleus of that atom. When the atoms get closer, the electrons are attracted by both nuclei.*



Formation of the covalent bond between hydrogen atoms.

While a nucleus attracts the electrons of another atom there are also repulsions both between the electrons and the nuclei of the atoms. When the attractive forces are greater than the repulsive forces the atoms get closer. When the attractive and repulsive forces become equal the electrons start to rotate around both nuclei (not only around the nucleus of their atom) and a bond is formed.

Depending upon the numbers of electrons that are shared, double and triple bonds may be formed. For example to be stable the oxygen atom needs two more electrons and so it forms a double bond with another oxygen atom ($O=O$). The nitrogen atom needs to gain three electrons to reach stability and so it forms a triple bond with another nitrogen atom, ($N\equiv N$).



Covalent bonds can be classified into three groups; nonpolar, polar and coordinate covalent bonds.

Nonpolar Covalent Bonds

These are bonds that are formed between two atoms with the same electronegativity values. In this kind of covalent bond, the attractive forces between both atoms and the bonding electrons are equal so the bond is nonpolar, meaning that the bonding electrons are shared equally between both atoms.

For example;

The bonds in H_2 molecule ($H-H$), N_2 molecule ($N\equiv N$), Cl_2 molecule ($Cl-Cl$) and O_2 molecule ($O=O$) are all nonpolar covalent bonds.

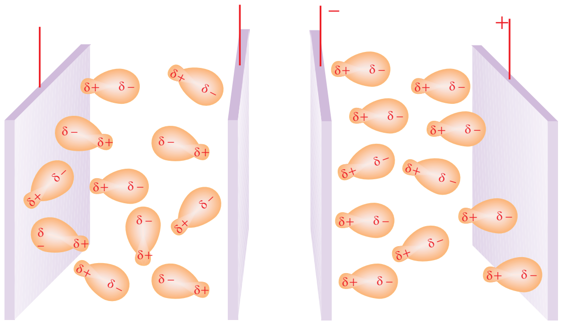
In these molecules the electronegativity difference between the atoms which form the bond is zero and therefore the charge distribution within the bond is equal.

Polar Covalent Bonds

As the electronegativity difference between the atoms increases, the attraction of the nuclei for the bonding electrons starts to differ. The atom with the greater electronegativity value attracts the bonding electrons more. But this increased attractive force is not so great as to completely take the bonding electrons and form an ion. In these covalent bonds, since the atom having the greater electronegativity value has more attraction for electrons, the bonding electrons are not shared equally. Thus electron density is not distributed equally between the atoms and the covalent bond has partially positive and partially negative poles. This bond is called a ***polar covalent bond***. In polar covalent bonds, the electron density distribution depends upon the electronegativities of the atoms. For example, let's examine the covalent bond between the $H\cdot$ and $\cdot\overset{\cdot\cdot}{\underset{\cdot\cdot}{Cl}}:$ atoms.

In the HCl molecule, the shared electrons are attracted more by the chlorine atom (electronegativity value 3.0) than by the hydrogen atom (electronegativity value 2.1). But the total transfer of electrons from hydrogen to chlorine does not happen because the electronegativity difference between hydrogen and chlorine is only 0.9, less than the 1.9 needed to form an ionic bond.

So between hydrogen and chlorine, a polar covalent bond having unequal charge distribution forms. In this molecule, the chlorine end of HCl molecule becomes partially negative, and the hydrogen end becomes partially positive. However, the negative charge is equal to the positive charge and the molecule overall is neutral.



neutral field

electrically charged field

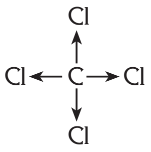
Charges in polar molecules move to the opposite sides in an electrical field.

Most chemical bonds are neither totally covalent nor totally ionic. As the difference in electronegativities between the two atoms increases, chemical bonds change from nonpolar covalent to polar covalent and then to ionic as the polarity of the bond increases.

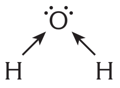
EXTRA

Polarity of Molecules

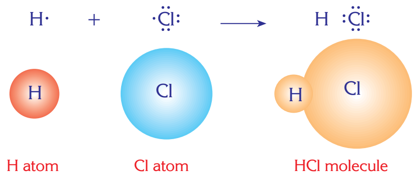
In a polar covalently bonded compound, the overall molecule might be polar or non-polar depending on the geometry of the molecule. Consider CCl_4 and H_2O. In both compounds, the elements possess different electronegativities so the bonds are polar.



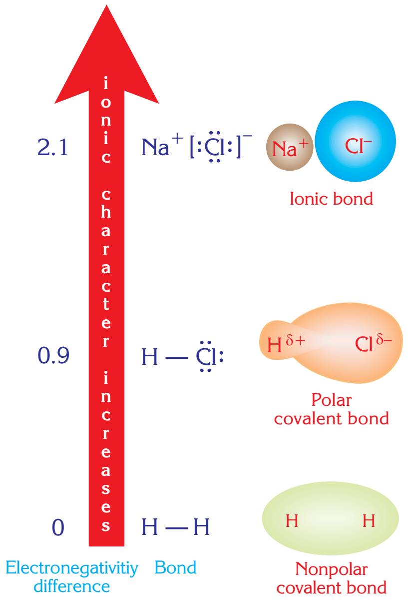
When we look at the overall molecular structure of carbon tetrachloride, the net vectorial force in this molecule is zero as its shape is symmetrical so CCl_4 is a non-polar molecule.



But in the water molecule the polar forces do not cancel one another out therefore the molecule is polar.



Formation of the HCl molecule by hydrogen and chlorine atoms.



The ionic character of a bond increases with increasing electronegativity difference between the bonding elements.

Coordinate Covalent Bonds

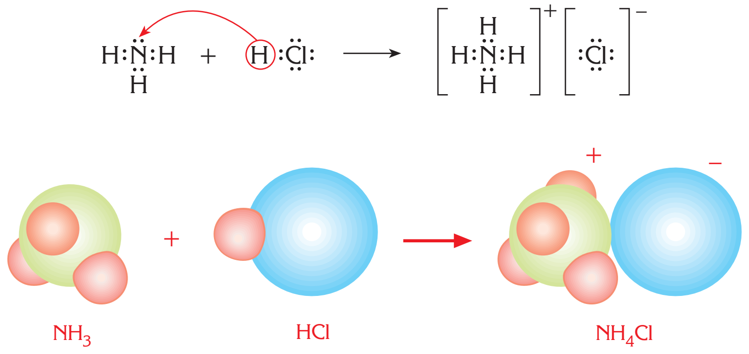
In the formation of certain compounds a covalent bond can be formed in which both of the shared electrons come from only one of the atoms. These bonds are called ***coordinate covalent bonds***. Let's examine the formation and bond structure of the NH_4^+ ion which contains a coordinate covalent bond.

$NH_3 + HCl \rightarrow NH_4Cl$

The Lewis dot structures of NH_3 and HCl are $\mathrm{H}:\underset{\mathrm{H}}{\underset{\cdot\cdot}{\overset{\cdot\cdot}{\mathrm{N}}}}:\mathrm{H}$ and $\mathrm{H}:\underset{\cdot\cdot}{\overset{\cdot\cdot}{\mathrm{Cl}}}:$

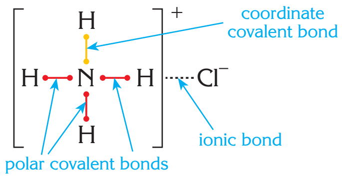
The hydrogen atom within a HCl molecule has shared its valence electron with the chlorine atom which has the greater attraction for the bonding electrons. This causes the HCl bond to be polar, with the hydrogen atom having a partial positive charge. This hydrogen is then attracted towards the lone (unshared) electron pair on the NH_3 molecule to form a covalent bond. The HCl bond breaks, leaving chlorine with both the bonding electrons.

In this new N — H bond both of the shared (bonding) electrons come from nitrogen.



Formation of the coordinate covalent bond.

In the NH_4^+ ion the N:H coordinate covalent bond is formed from the donation of an unshared electron pair while the other three (N:H) bonds are polar covalent bonds. Once it has been formed there is no difference between a coordinate covalent bond and other bonds. In other words, all the N — H bonds in the NH_4^+ ion are the same.

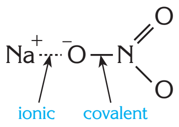


In the NH_4Cl structure while the NH_4^+ ion contains four covalent bonds (1 coordinate covalent bond and 3 polar covalent bonds): there is an ionic bond between the NH_4^+ and Cl^- ions.

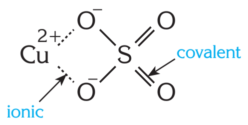
EXTRA

Both Ionic and Covalent

Some molecules contain both ionic and covalent bonds. For example we can draw the molecular structure of $NaNO_3$ and $CuSO_4$, as



and



Example 4

Compare the polarity and ionic character of the bonds formed between the following pairs. (Refer to Figure 1)

H—I, Si—F, N—H, Rb—F

Solution

Let's find the electronegativity values of the elements by using figure 1 and then calculate the electronegativity differences.

I—H	F—Si	N—H	F—Rb
$2.5 - 2.1 = 0.4$	$4.0 - 1.8 = 2.2$	$3.0 - 2.1 = 0.9$	$4.0 - 0.8 = 3.2$

As we know, the polarity of the bonds depends upon the electronegativity differences.

Let's arrange the bonds according to their electronegativity difference;

Bond :	H—I	N—H	Si—F	Rb—F
Electronegativity difference :	0.4	0.9	2.2	3.2

Here the H—I bond has the lowest electronegativity difference while the Rb—F bond has the highest. So amongst these bonds, the H—I molecule is the least polar and has the least ionic character, and the Rb—F bond is the most polar and has the most ionic character.

The increasing order of the ionic character of these bonds are;

H—I < N—H < Si—F < Rb—F

Example 5

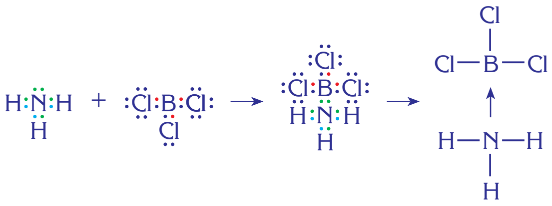
Identify the types of the bonds in the NH_3BCl_3 molecule which is formed by the reaction $NH_3 + BCl_3 \rightarrow$ ....

Solution

The electron dot structures of NH_3 and BCl_3 molecules are

respectively.

So there are 3 (N — H) polar covalent bonds in the ammonia (NH_3) molecule and 3 (B — Cl) polar covalent bonds in the boron trichloride (BCl_3) molecule.



Boron which is the central atom in boron trichloride has no unshared electrons. Therefore it cannot donate electrons to form a bond. But nitrogen (N) in the ammonia molecule has 1 unshared electron pair and it can therefore form a coordinate covalent bond with boron.

As a result, the NH_3BCl_3 molecule contains 6 polar covalent bonds

[3 (N — H) and 3 (B — Cl)]

and 1 coordinate covalent bond (N–B), so in total there are 7 covalent bonds.

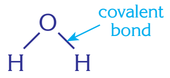
Example 6

What kinds of chemical bonds do the following compounds contain?

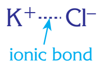
Explain briefly. **a.** H_2O **b.** KCl **c.** Na_3PO_4

Solution

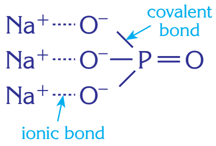
a. In the H_2O molecule, between the hydrogen and oxygen atoms (both nonmetals) there are polar covalent (O — H) bonds.



b. In the KCl structure, potassium (K) is a metal and chlorine (Cl) is a nonmetal. So there is an ionic bond between K and Cl due to their high electronegativity differences.



c. In the Na_3PO_4 compound between the three Na^+ ions and the PO_4^{3-} ion there are ionic bonds. In the structure of the PO_4^{3-} ion there are polar covalent bonds which are formed between the P and O nonmetal atoms.



As a result, the Na_3PO_4 compound contains both ionic and polar covalent bonds.

3. HYBRIDIZATION

The mixing of different orbitals which have closer energy levels, to form new orbitals with the same energy level is called ***hybridization***. The new orbitals formed at this new energy level are called hybrid orbitals.

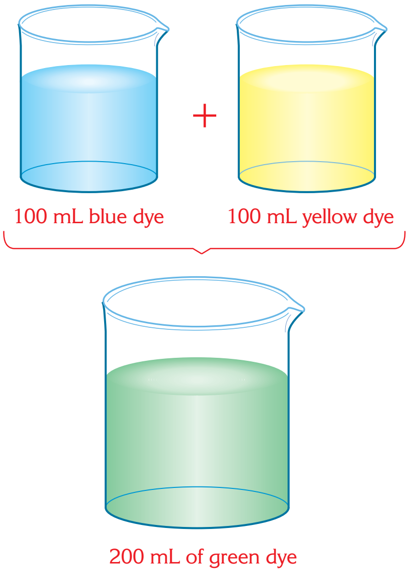
Hybridization occurs between two or more different types of orbitals (generally s, p or d orbitals). For example, there are three types of hybrid orbitals which may occur between the s and p orbitals, these are named as sp, sp^2 and sp^3 hybrid orbitals.

It is not possible to form hybrid orbitals between the same type of orbital. For example s orbitals cannot form ss hybrid orbitals and p orbitals cannot form pp hybrid orbitals.

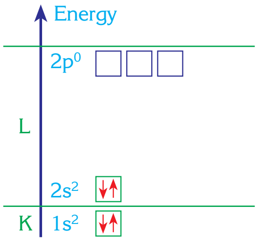
Group 2A elements of the periodic table can undergo sp hybridization, group 3A elements can undergo sp^2 hybridization and group 4A elements can undergo sp^3 hybridization. Molecules formed by atoms of these groups generally contain bonds with hybridized orbitals. Since hybrid orbitals overlap with each other, stable molecules are formed. Maximum overlapping often occurs between the hybrid orbital of one atom and the orbital of another atom, molecules formed in this way have lower energies. The energy needed for hybridization is balanced against the energy which is released during bond formation.

Hybridization occurs during the formation of a chemical bond. It is not possible to occur in an individual atom. Hybrid orbitals play an important role in determining the geometric shape of a molecule.

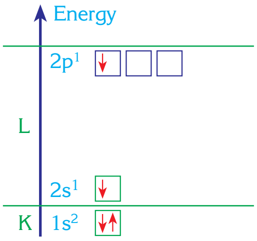
Now let's study sp, sp^2 and sp^3 hybridization in detail.



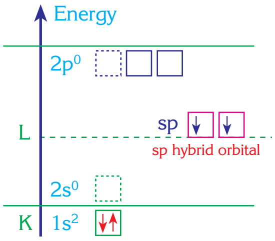
When 100 mL of blue dye is mixed with 100 mL of yellow dye 200 mL of green dye is formed. At the same way, when s and p orbitals are mixed (hybridized), hybrid orbitals which have both the characteristics of s and p orbitals are formed.



a. *Beryllium atom in its ground state level*



b. *Beryllium atom in an excited state*



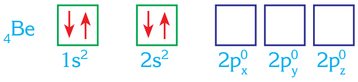
c. *sp hybridization in the beryllium atom*

Figure 4: *Energy changes during the formation of sp hybrid orbitals in the beryllium atom.*

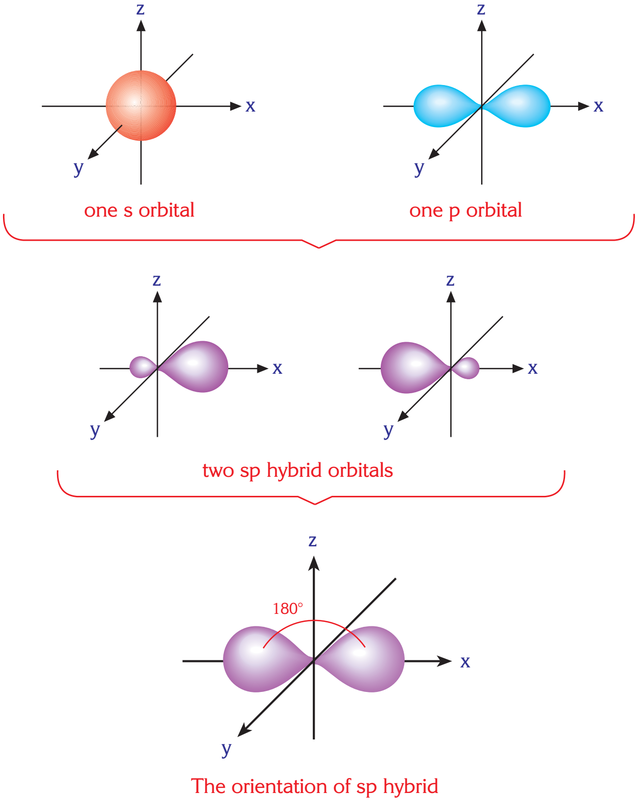
A model of the sp hybrid orbital

3.1. sp HYBRIDIZATION

Let's look at the ground state electron configuration and orbital diagram of the beryllium atom ($_4$Be) which is the first element in group 2A.



As it does not have any unshared electrons, beryllium would not be expected to form a covalent bond. But experimentally it is found that beryllium is able to form two covalent bonds. To form these bonds one electron moves from the 2s orbital to the 2p orbital leaving the atom in an excited state with two unpaired electrons (Figure 4b).



The orientation of sp hybrid orbitals is linear

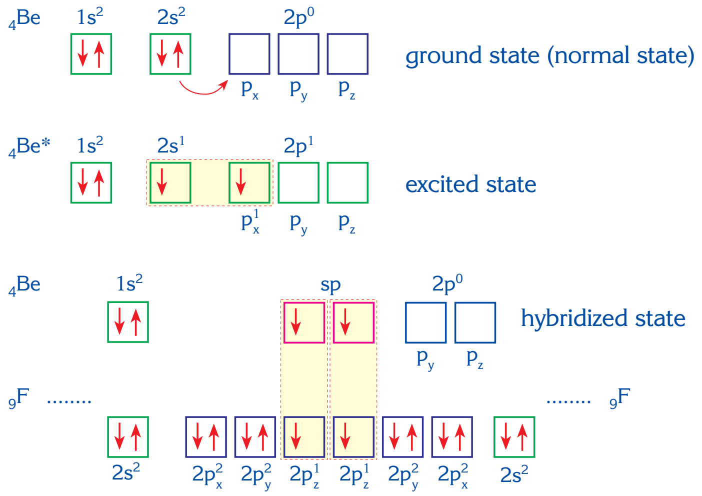
Two sp hybrid orbitals are formed as a result of mixing one s orbital with one p orbital. The energy of the sp hybrid orbitals is greater than the s orbital but less than the p orbitals. Each sp orbital has 50% s character and 50% p character.

Example 7

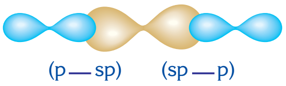
Show the hybridization of the beryllium atom when it bonds with fluorine.

$_{4}Be$, $_{9}F$

Solution



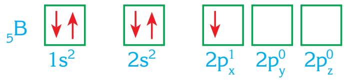
As it is shown, Be forms two bonds with two different F atoms:



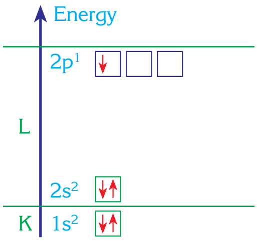
In F — Be — F, the two bonds are formed by the overlap of sp orbitals with p orbitals.

3.2. sp^2 HYBRIDIZATION

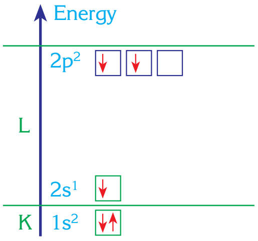
Let's look at the ground state electron configuration and orbital diagram of Boron ($_{5}B$) which is the first element of group 3A.



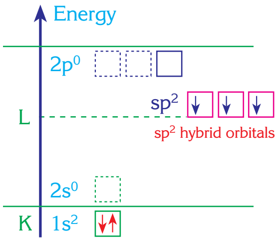
It is found experimentally that boron can form three covalent bonds. But as it has only one unpaired valence electron in the ground state, it appears only to be able to form one bond. To create three unpaired electrons, one electron in the 2s orbital is promoted to the $2p_y$ orbital. To form three identical bonds with the same energy, two p and one s orbitals mix to give three sp^2 orbitals. These three identical and half filled sp^2 orbitals enable boron to form three identical bonds.



a. *Boron atom in its ground state level.*



b. *Boron atom in an excited state.*

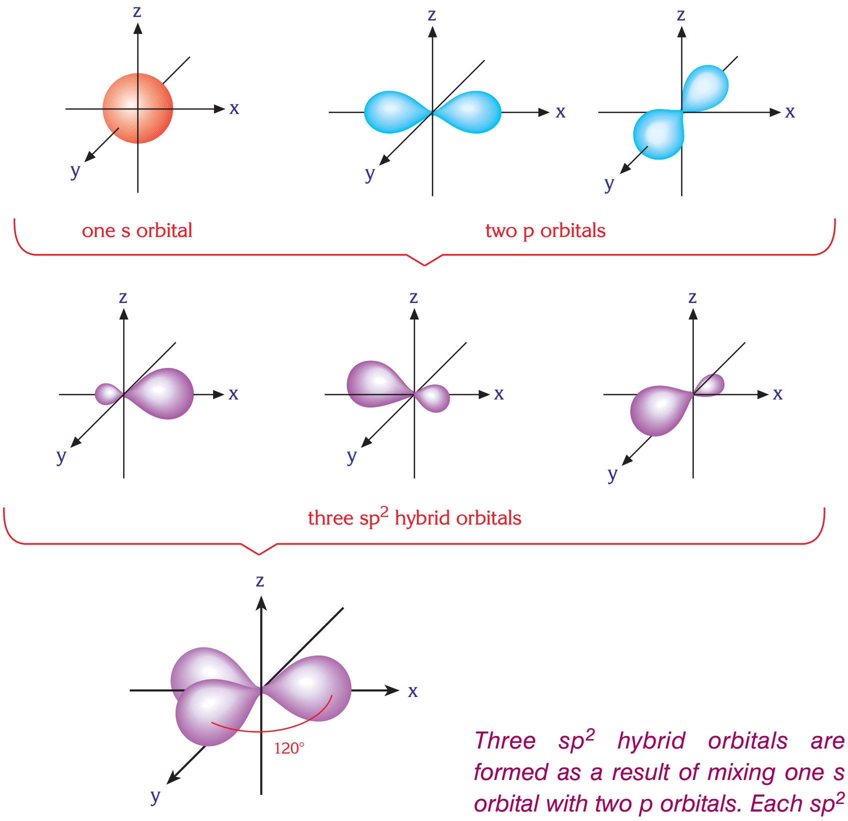


c. *sp^2 hybridization in the boron atom.*

Energy changes during the formation of sp^2 hybrid orbitals in the boron atom.

A model of sp^2 hybrid orbitals

In the excited state of boron, one of the valence electrons is in the s and the other two electrons are in the p orbitals. So if hybridization did not occur, the three bonds that would form would have different lengths and different properties.



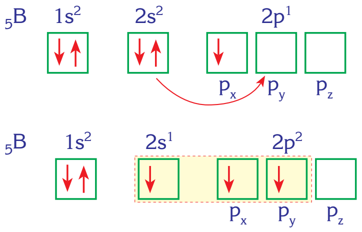
The orientation of three sp^2 hybrid orbitals is trigonal planar

Three sp^2 hybrid orbitals are formed as a result of mixing one s orbital with two p orbitals. Each sp^2 hybrid orbital has 33.3% s and 66,7% p character.

Example 8

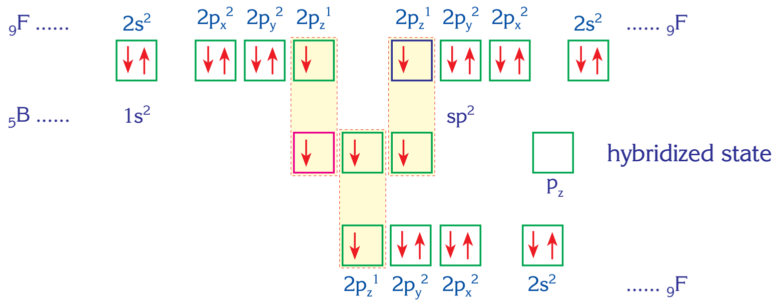
Show the hybridization of the boron atom when it bonds with fluorine. $_5B$, $_9F$

Solution

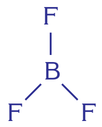


ground state (normal state)

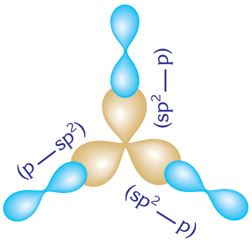
excited state



So boron forms three bonds with three fluorine atoms:

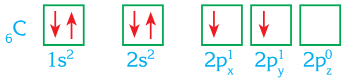


there is overlap of the sp^2 hybrid orbitals with the p orbitals of F.



3.3. sp^3 HYBRIDIZATION

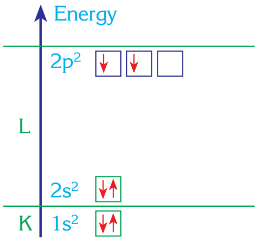
Lets look at the ground state electron configuration and orbital diagram of carbon ($_6$C) which is the first element in group 4A.



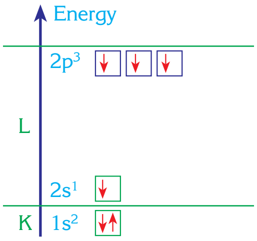
In this case since carbon has only two unpaired electrons, it seems likely that it will only form only two covalent bonds, but it is known that carbon can form four covalent bonds. To form four bonds, one electron is promoted from the 2s orbital to the $2p_z$ orbital. Then the one 2s orbital and three 2p orbitals mix together to form four new sp^3 hybrid orbitals as shown in Figure 5. So in this case of hybridization, three p and one s orbital combine to give four identical sp^3 orbitals.

The carbon atom can also undergo sp^2 and sp hybridization. Later we will study the sp and sp^2 hybridization of carbon when it forms double and triple bonds.

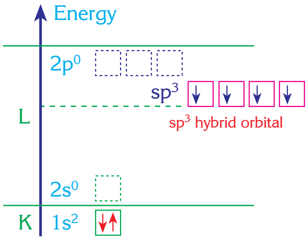
A model of sp^3 hybrid orbitals.



a. *Carbon atom in its ground state*

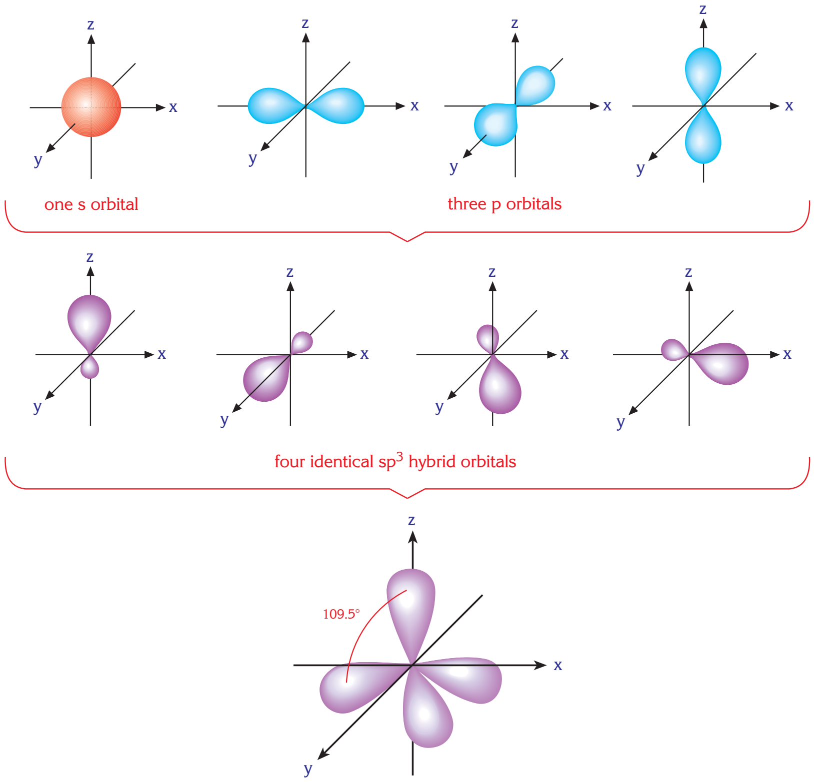


b. *Carbon atom in an excited state*



c. *sp^3 hybridization in the carbon atom*

Figure 5: *Energy changes during the formation of the sp^3 hybrid orbitals in a carbon atom.*



The orientation of four sp^3 hybrid orbitals is tetrahedral

Four sp^3 hybrid orbitals are formed as a result of mixing one s orbital with three p orbitals. Each sp^3 hybrid orbital has 25% s and 75% p character.

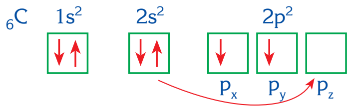
Example 9

Show the kinds of hybridization when the carbon atom bonds with chlorine.

${}_{6}C$, ${}_{17}Cl$

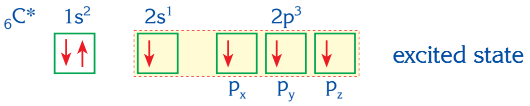
Solution

Electron configuration of the carbon atom:

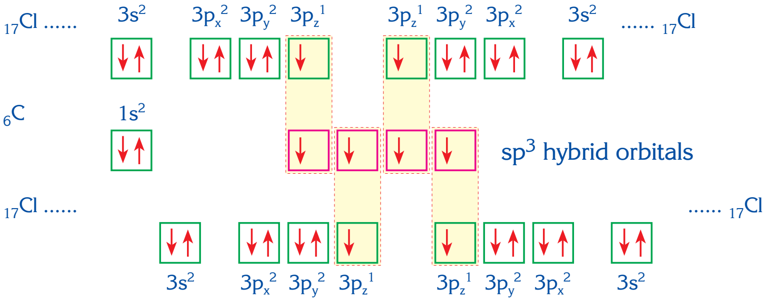


ground state (normal state)

when it is excited:

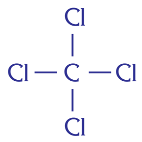


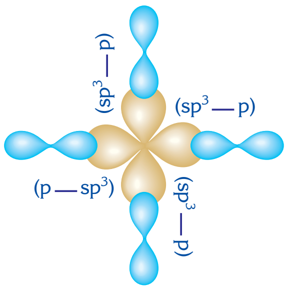
if 2s and three 2p orbitals mix to form four hybrid orbitals, they can overlap with the unpaired p electrons in four different chlorine atoms:



So the carbon atom forms four bonds with four chlorine atoms.

The molecular structure is: The overlap of orbitals





4. COVALENT BONDING CAPACITY OF THE SECOND ROW ELEMENTS

The number of covalent bonds that an element can form is equal to the number of unpaired valence electrons of that element.

Therefore the number of half-filled orbitals indicates the number of bonds that the atoms can form. Elements in the same group of the periodic table exhibit similar chemical properties as they have the same number of valence electrons. We will explain bond formation of one representative element from each main group. The other elements found in the same group generally form bonds in a similar way.

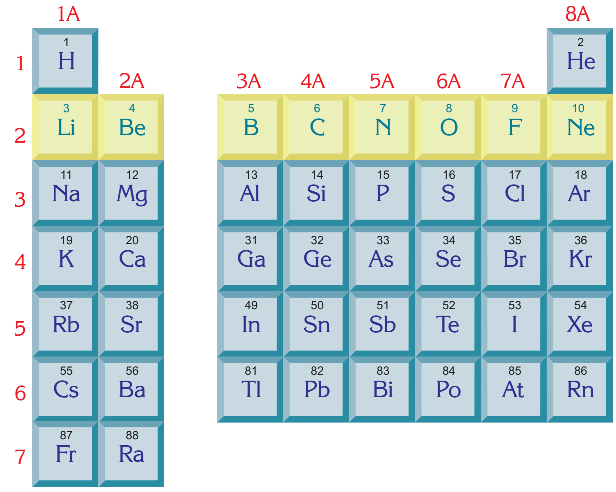
EXTRA

The VSEPR (Valence shell electron pair repulsion) Model

Atom are bonded to each other in molecules by the sharing of pairs of valence shell electrons. But electron pairs repel one another. Therefore, electron pairs try to stay out of each other's way as far as possible.

The best arrangement of a given number of electron pairs is the one that minimizes the repulsion among them. This simple idea is the basis of the VSEPR. This model is used to predict shapes of molecules.

The second row of the periodic table consists of lithium (Li), beryllium (Be), boron (B), carbon (C), nitrogen (N), oxygen (O), fluorine (F) and neon (Ne). Now let's examine the compounds of these elements form with hydrogen.



	1A	2A	3A	4A	5A	6A	7A	8A
1	1 H							2 He
2	3 Li	4 Be	5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg	13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	19 K	20 Ca	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
5	37 Rb	38 Sr	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
6	55 Cs	56 Ba	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
7	87 Fr	88 Ra						

Second row elements of the periodic table.

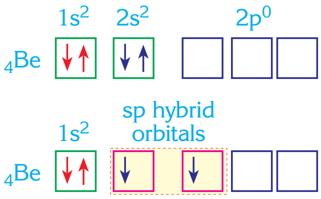
4.1. BONDING CAPACITY OF LITHIUM

Lithium is a metal so it tends to form an ionic bonds with non-metals. The compound lithium hydride, LiH, is made up of crystals with a cubic lattice structure.

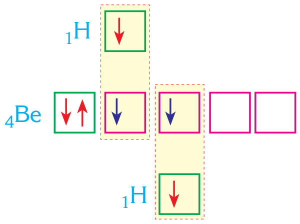
4.2. BONDING CAPACITY OF BERYLLIUM

The electron configuration of Be is $1s^2$ $2s^2$, it has two valence electrons in its ground state. It should not be able to form a covalent bond as the electrons are paired.

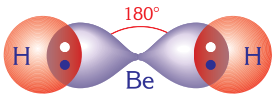
To form a bond, the filled 2s orbital and one of the 2p orbitals combine and give two half-filled sp orbitals.



Therefore beryllium can have two half filled orbitals and two (unpaired electrons) in its excited state and form the BeH_2 molecule with hydrogen. The bond structure of BeH_2 is given below;



$H{:}Be{:}H \longrightarrow H-Be-H \longrightarrow BeH_2$



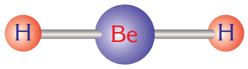
BeH_2 molecule. The direction of orbitals is linear.

The BeH_2 molecule is formed between Be and H atoms.

Because of the electronegativity difference (0.6) between Be (1.5) and H (2.1), the Be—H bonds are polar.

The geometry of the molecule

The two hydrogen atoms having the same electronegativity value cause the BeH_2 molecule to be ***nonpolar***. This is because the molecule is symmetrical and the net vectorial force applied on Be atom by the bond dipoles is zero.

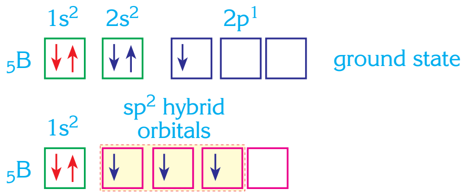


The shape of the BeH_2 molecule.

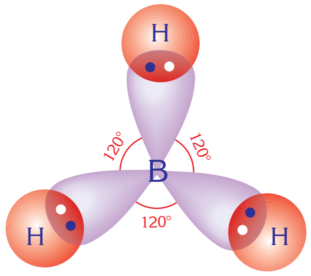
4.3. BONDING CAPACITY OF BORON

Although the boron atom (with electron configuration $1s^2\ 2s^2\ 2p^1$) has three valence electrons, only one of them is unpaired in the ground state.

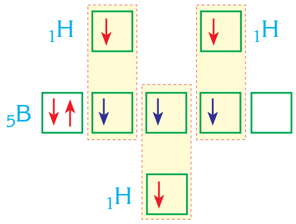
To increase the number of unpaired electrons, one electron is promoted from the 2s orbital to a 2p orbital. Then the 2s and two 2p orbitals mix to form three identical sp^2 hybrid orbitals.



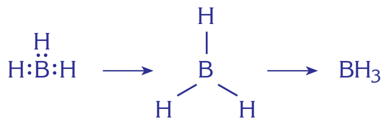
As a result of hybridization, boron can form three bonds and so the BH_3 molecule is formed with hydrogen.

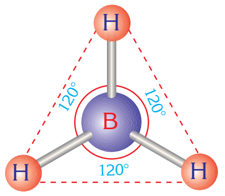


In the BH_3 molecule the orientation of orbitals is trigonal planar.



B forms the BH_3 molecule with H.





The shape of the BH_3 molecule.

Because of the electronegativity difference (0.7) between B(2.8) and H(2.1), B — H bonds are polar.

The geometry of the molecule

The geometry of the BH_3 molecule is ***trigonal*** planar. The net vectorial force applied on the boron atom by the three polar bonds is zero due to the symmetrical shape, so the molecule is ***nonpolar***.

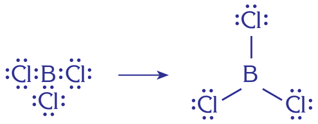
Many of the compounds of the other elements in group 3A have similar bond structure and geometry to the BH_3 molecule.

Example **10**

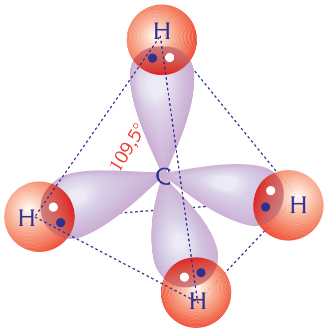
Explain the bond structure of the BCl_3 molecule by using electron dot structure. ($_5B$, $_{17}Cl$)

Solution

Boron undergoes sp^2 hybridization and forms three identical sp^2 hybrid orbitals containing three unpaired electrons. Chlorine has 7 valence electrons of which just one of them is unpaired. Unpaired electrons are shared by three chlorine atoms and a boron atom to form BCl_3.

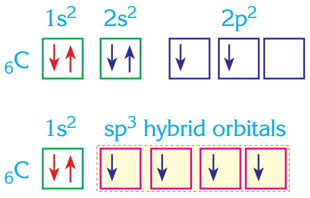


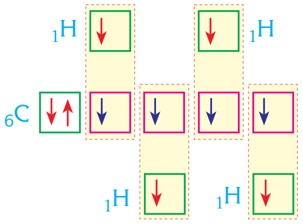
4.4. BONDING CAPACITY OF CARBON



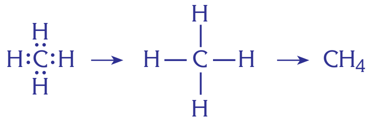
The orientation of orbitals in CH_4 molecule is tetrahedral.

The carbon atom ($_6C$) has the electron configuration of $1s^22s^22p^2$. There are 4 valence electrons, of which only two are unpaired in the ground state. During the formation of carbon compounds, one 2s and three 2p orbitals combine to give four identical sp^3 orbitals by the promotion of an electron from the 2s orbital to a 2p orbital. These 4 unpaired orbitals then mix to form four identical sp^3 hybrid orbitals.





Between C and H the CH_4 molecule is formed.



As a result of this hybridization, carbon forms four bonds with hydrogen to form the CH_4 molecule.

The geometry of the molecule

The shape of the CH_4 molecule is ***tetrahedral***. A tetrahedral orientation of equal bonds (which are formed from the overlap of the identical sp^3 hybrid orbitals and the hydrogen 1s orbitals) gives a bond angle of 109.5° (Figure 6).

In the CH_4 molecule the net vectorial force applied on carbon atom by the four polar bonds is zero. This is because of the symmetry of the molecule, hence it is ***non–polar***.

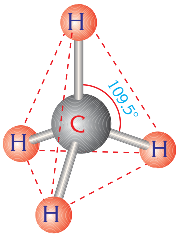
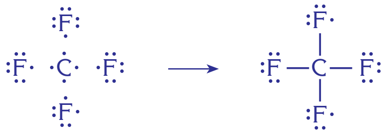


Figure 6: *The shape of the CH_4 molecule.*

Example **11**

Explain the bond structure of the CF_4 molecule by using electron dot structure.

Solution

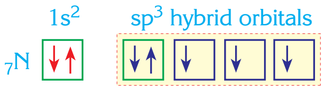


Carbon undergoes hybridization and forms four identical sp^3 hybrid orbitals. Only one of seven valence electrons in fluorine is unpaired. Four identical half filled sp^3 hybrid orbitals of carbon are filled with the four unpaired electrons of four fluorine atoms.

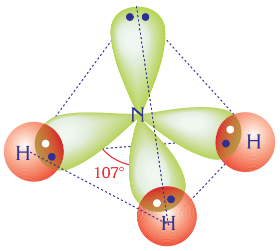
In CF_4 the bonds between carbon and fluorine are polar. The shape of the molecule is tetrahedral hence the attractive forces of the four dipoles (one for each polar bond) cancel each other out. Therefore the molecule is non-polar.

4.5. BONDING CAPACITY OF NITROGEN

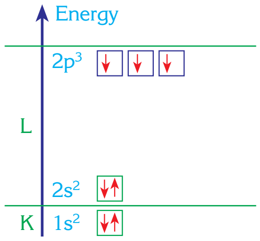
The electron configuration of nitrogen $1s^22s^22p^3$ shows that there are five valence electrons. Three of them are unpaired in this state so nitrogen can form three bonds, however, hybridization still occurs, with the s and p orbitals mixing to form four sp^3 hybrid orbitals.



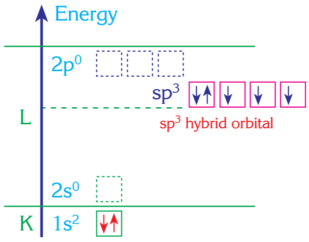
If nitrogen uses only its p orbitals in bond formation, the angle between N–H bonds would be 90°. However, compounds prefer formations in which electrons are as far apart as possible. For ammonia this is made possible by forming a tetrahedral structure in which the angle between the bonds (N–H) is 107°. This is only possible by undergoing sp^3 hybridization.



The orientation of the orbitals in the NH_3 molecule is trigonal pyramidal.



a. *The ground state energy level of the nitrogen atom*



b. *sp^3 hybridization in the nitrogen atom*

Figure 7: *Formation of sp^3 hybrid orbitals in the nitrogen atom.*

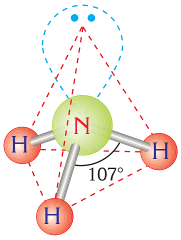
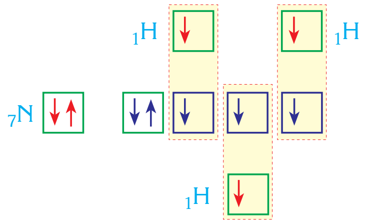
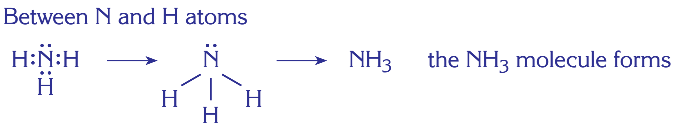


Figure 8: *The shape of the NH_3 molecule is trigonal pyramidal.*

Nitrogen forms three bonds by using its three half-filled sp^3 hybrid orbitals to form the NH_3 molecule with hydrogen as shown in Figure 7.





Because of the electronegativity difference (0.9) between N(3.0) and H(2.1), N — H bonds are polar.

The geometry of the molecule

In ammonia, three of the five sp^3 electrons take part in the bond formation of N — H bonds.

The other two electrons are found around the nitrogen atom as a free electron pair. Since the free electron pair does not form a bond, it occupies a larger volume in space then the bonding electron pairs between the nitrogen and hydrogen atoms.

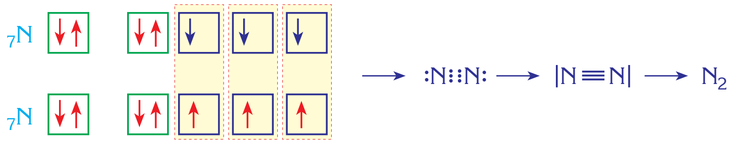
Due to the greater repulsive effect of the free electron pair in the NH_3 molecule, the N — H bonds get pushed together slightly. So the molecular geometry of NH_3 molecule is different from that of the BH_3 molecule in that it is ***trigonal pyramidal***. Unlike the ***non-polar*** BH_3, the NH_3 molecule is ***polar***. The angles between the nitrogen – hydrogen bonds are 107° (Figure 8) in NH_3.

Example 12

Show the bond structure of the nitrogen molecule by using an orbital diagram, electron dot structure and line representation.

Solution

Nitrogen has one filled and three half filled valence orbitals. Two nitrogen atoms form three bonds with their three half-filled orbitals. The remaining free pairs of electrons (one on each N atom) are placed around the nitrogen atoms.

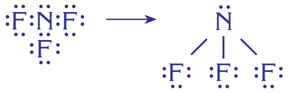


Example 13

Explain the bonding and molecular structure of the NF_3 molecule by using electron dot representation. ($_7N$, $_9F$)

Solution

The valence electrons of nitrogen and fluorine are five and seven respectively. In this case, to complete its octet nitrogen needs three more electrons and fluorine needs one more electron. Therefore one nitrogen atom combines with three fluorine atoms.

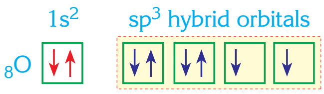


The N—F bonds are polar covalent. Due to the repulsive force of the lone pair electrons on the nitrogen atom the shape of the NF_3 molecule is trigonal pyramidal. The dipole forces do not cancel each other out so the molecule is ***polar***.

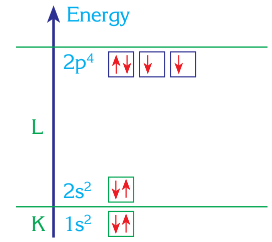
4.6. BONDING CAPACITY OF OXYGEN

Oxygen ($_8O$) has six valence electrons. Two of them are unpaired and the others are paired when the atom is in its ground state. However, advanced studies have shown that all four valence orbitals of oxygen are identical so when oxygen reacts with another element it combines its one 2s and three 2p orbitals to form four identical sp^3 orbitals. Two of the six valence electrons of oxygen take part in bond formation.

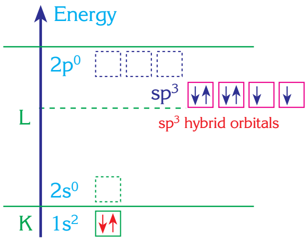
As it was mentioned in the formation of the NH_3 molecule, compounds prefer configurations in which the electron pairs are as far apart as possible. Therefore oxygen undergoes sp^3 hybridization resulting in a tetrahedral shape.



Oxygen forms two bonds by utilising its half filled sp^3 hybrid orbitals when it forms the H_2O molecule with hydrogen.

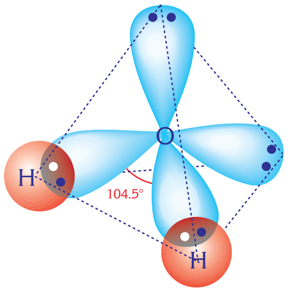


***a.** Energy levels of the oxygen atom in its ground state.*

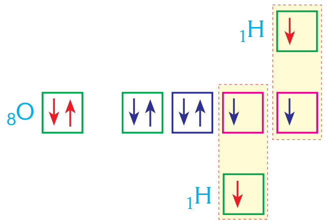


***b.** sp^3 hybridization in the oxygen atom.*

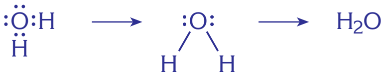
Formation of sp^3 hybrid orbitals in the oxygen atom



The orientation of the orbitals in the H_2O molecule.



Between O and H atoms



the H_2O molecule forms.

Because of the electronegativity difference (1.4) between O (3.5) and H (2.1), O — H bonds are polar.

The geometry of the molecule

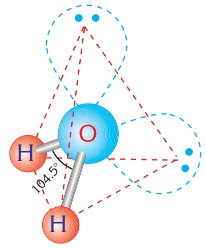


Figure 9 : *The shape of the H_2O molecule*

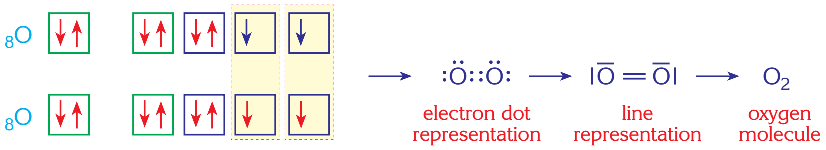
The two unpaired electrons of oxygen atom form two polar O — H bonds in a water molecule. The other four valence electrons around the oxygen atom exist as two free electron pairs. Since the free electron pairs of oxygen do not form bonds, they occupy a larger volume than bonding electron pairs. Due to the greater repulsive effect of the free electron pairs compared with the bonding electrons in the H_2O molecule, the shape of the H_2O molecule is ***bent***. The angle between oxygen–hydrogen bonds is 104.5° (Figure 9) and the molecule is ***polar***.

Example 14

Show the bond structure of the oxygen molecule by using orbital, electron dot and line representations.

Solution

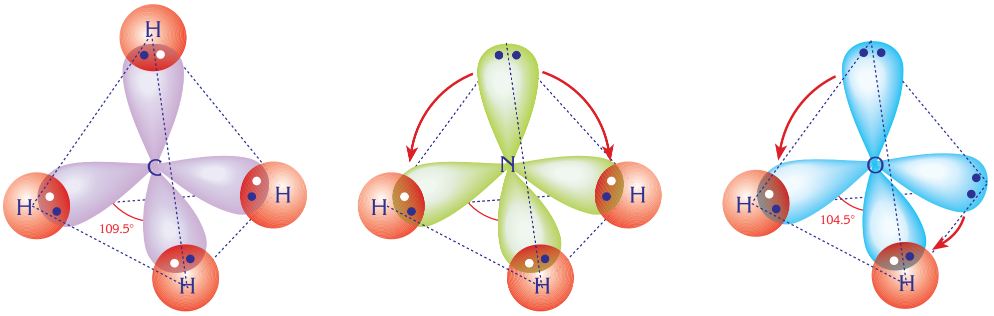
Oxygen has two filled, and two half filled valence orbitals. The oxygen atom forms two bonds by overlapping its half-filled orbitals with another oxygen atom.



orbital representation

EXTRA

Both Ionic and Covalent



Molecule formula :	CH_4	NH_3	H_2O
Molecule geometry :	tetrahedral	trigonal pyramidal	bent
Polarity of the molecule :	nonpolar	polar	polar
Bond angle :	$H\widehat{C}H$: 109.5°	$H\widehat{N}H$: 107°	$H\widehat{O}H$: 104.5°

The shape of methane (CH_4), ammonia (NH_3) and water (H_2O) molecules are based upon the tetrahedron. CH_4 has four bonds while the other two molecules have lone pair electrons as well as bonding electron pairs. As the number of lone pair electrons increases the bond angles decrease. The reason is that the free electron pairs occupy a larger volume in space. As a result, the shape of CH_4 is tetrahedral, but due to the greater repulsive forces of free electron pairs the shape of NH_3 is trigonal pyramidal with a bond angle of 107° and that of H_2O is bent with a bond angle of 104.5°.

To summarize, the orientations of the electron pairs around of the central atoms in each of the three molecules are based upon the tetrahedron.. The shape of methane is tetrahedral, but in the ammonia and water molecules due to the repulsive forces of the non-bonding electrons, the shapes are trigonal pyramidal and bent respectively, with a decreasing bond angle.

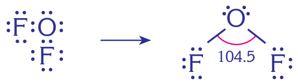
Example 15

Show the bond structure of the OF_2 molecule by using electron dot representation.

Solution

The number of valence electrons of oxygen and fluorine are six and seven respectively. So oxygen needs to share two electrons and fluorine one electron to complete its octet. Therefore one oxygen atom combines with two fluorine atoms.

The shape of OF_2 is bent. Both the (O — F) bonds and the OF_2 molecule are polar.

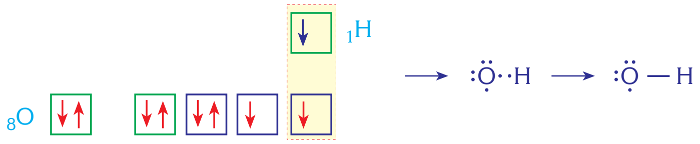


Example 16

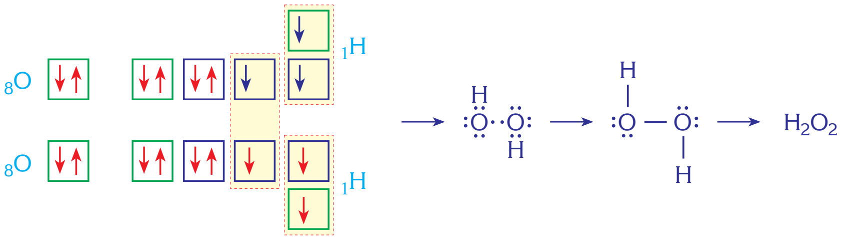
Show the bond structure of the H_2O_2 molecule by using orbital diagrams, electron dot structure and line representation.

Solution

When one hydrogen atom is attached to one oxygen atom OH forms



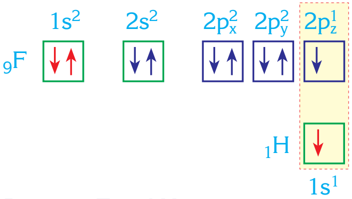
But in this case OH has one more half filled orbital and it is very reactive. To achieve stability, the half filled orbital of OH overlaps with the half filled orbital of another OH and the H_2O_2 molecule is formed.



4.7. BONDING CAPACITY OF FLUORINE

Although fluorine has seven valence electrons, only one of them is unpaired, so the fluorine atom can form one bond. The formula of the compound formed between hydrogen and fluorine is HF and its bond structure is as follows;

Distribution of bonding electrons in the HF molecule.



Between F and H atoms

$H\ :\underset{\cdot\cdot}{\ddot{F}}: \longrightarrow H-\underset{\cdot\cdot}{\ddot{F}}: \longrightarrow HF$

the HF molecule is formed.

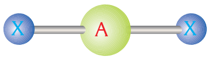
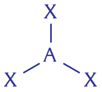
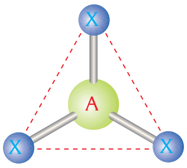
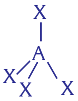
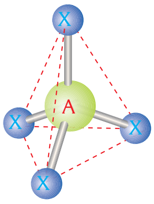
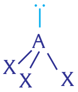
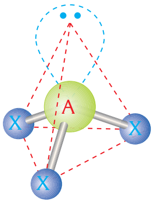
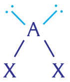
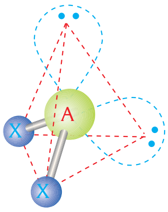
The electronegativity difference (1.9) between F(4.0) and H(2.1) is very high, therefore the H — F bond is very polar.

The geometry of the molecule

The HF molecule is ***linear*** and as fluorine is more electronegative than hydrogen, the bonding electrons are closer to the fluorine atom.

4.8. BONDING CAPACITY OF NEON

Neon has eight valence electrons and all of them are paired, hence the valence orbitals of neon are completely filled. Therefore neon is very unreactive and does not bond with any other element. Similarly, the group 8A elements (noble gases) helium and argon are very unreactive. However, krypton and xenon may form bonds under certain conditions.

Orbitals used in bond formation	Nonbonding electron pairs	General formula	Molecular geometry	Molecular shape	Bond angle	Polarity of molecule	Example	The name of example compounds
sp	—	AX_2	X — A — X Linear		180°	Nonpolar	$BeCl_2$ BeF_2	Beryllium chloride Beryllium fluoride
sp^2	—	AX_3	 Trigonal planar		120°	Nonpolar	BCl_3 BF_3	Boron trichloride Boron trifluoride
sp^3	—	AX_4	 Tetrahedral		109.5°	Nonpolar	CCl_4 CH_4	Carbon tetrachloride Methane
sp^3	1	AX_3E	 Trigonal pyramidal		107°	Polar	NH_3 PCl_3	Ammonia Phosphorous trichloride
sp^3	2	AX_2E_2	 Bent		104.5°	Polar	H_2O OF_2	Water Oxygen difluoride

A : Central atom, X : Atoms bonded to central atom, E : Non-bonding electron pairs

Table 2 : *Molecules formed using hybrid orbitals in their bond formation.*

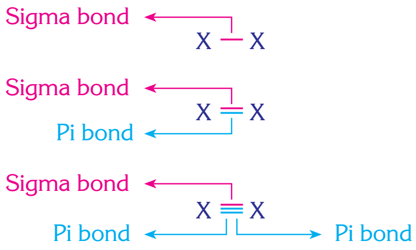
5. DOUBLE AND TRIPLE COVALENT BONDS

Some atoms, such as carbon, oxygen and nitrogen can form double or triple bonds as well as single bonds.

Two types of bonds may be formed when orbitals overlap. These are named sigma (σ) and pi (π) bonds.

All single bonds between two atoms are sigma (σ) bonds. Pi bonds can only be formed after a sigma bond has already been formed. Therefore a double bond contains one σ and one π bond, and a triple bond contains one σ and two π bonds.

Now let's examine the formation of σ and π bonds.

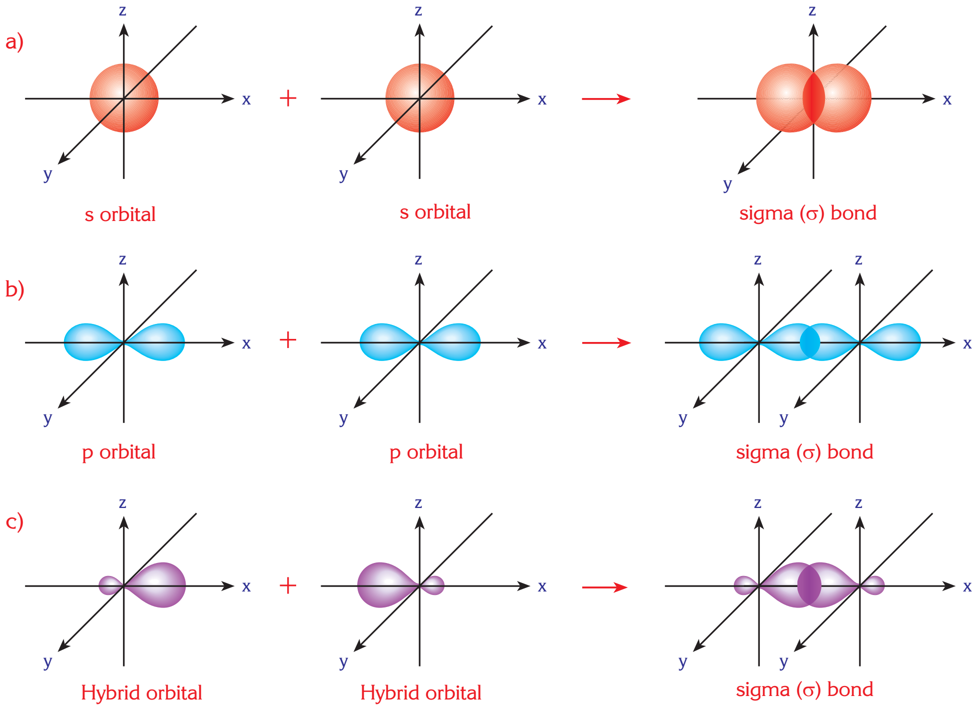


The distribution of σ and π bonds within single, double and triple bonds.

5.1. SIGMA (σ) BONDS

Sigma (σ) bonds are formed by the end to end overlap of two orbitals. This overlap can take place between s orbitals, p orbitals or hybrid orbitals.

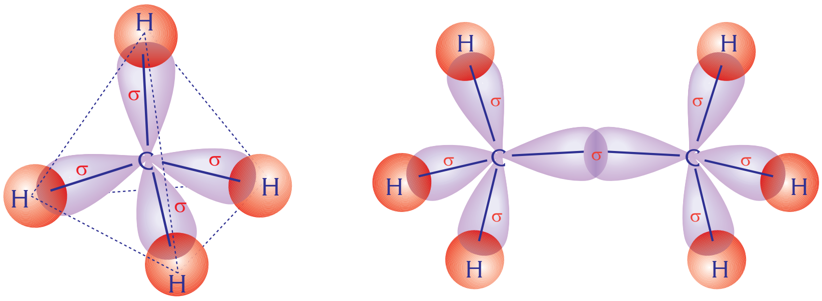
For example, in the methane molecule (CH_4), the four sp^3 hybrid orbitals of the carbon atom overlap end to end with one 1s orbital from each hydrogen atom to form four C — H bonds. Those bonds are all σ bonds.



Sigma bonds are formed from:

***a.** overlap of s orbitals, **b.** the end to end overlap of p orbitals, **c.** the overlap of hybrid orbitals*

Similarly, C — H sigma bonds in the C_2H_6 molecule are formed by the end to end overlap of sp^3 hybrid orbitals of the carbon atoms with the 1s orbitals of the hydrogen atoms. The C — C σ bond is formed by the end to end overlap of the sp^3 hybrid orbitals of the C atoms. So in the C_2H_6 molecule there are six C — H σ bonds and one C — C σ bond making seven σ bonds in total.



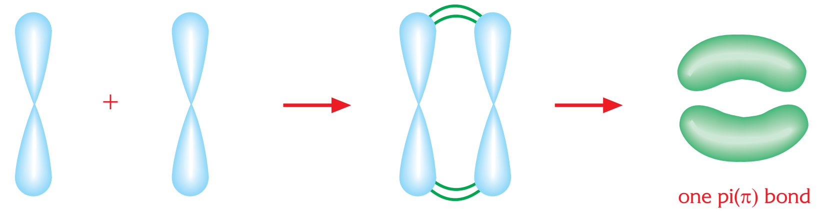
CH_4 molecule : 4 sigma(σ) bonds

C_2H_6 molecule : 7 sigma(σ) bonds

Sigma bonds in methane and ethane

5.2. Pi (π) BONDS

Pi (π) bonds are formed by the side by side overlap of two parallel p orbitals. In the π bond, the electron cloud lies above and below the plane formed by σ bonds. π bonds are weaker than σ bonds.



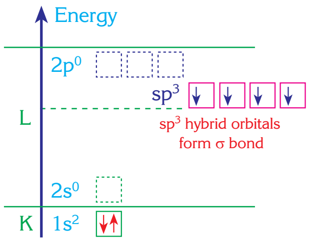
p orbital

p orbital

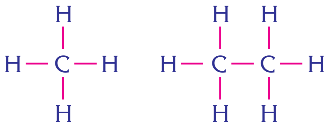
one pi(π) bond
(two separate
electron clouds)

A π bond is formed from the side by side overlap of unhybridized p orbitals.
p orbital + p orbital ⇒ 1 π bond (2 separate electron clouds)

A π bond can not be formed alone. It can be formed after the formation of a σ bond, if any unhybridized p–orbitals of atoms remain. In another words, to form a π bond, two atoms must form a σ bond first.



When carbon atoms undergo sp^3 hybridization, the hybrid orbitals form sigma bonds.



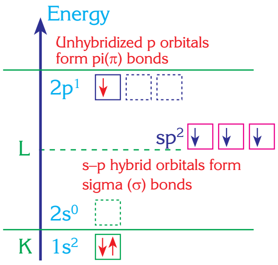
$$\begin{array}{ccccc} & & H & & \\ & & | & & \\ H & - & C & - & H \\ & & | & & \\ & & H & & \end{array} \qquad \begin{array}{ccccccc} & & H & & H & & \\ & & | & & | & & \\ H & - & C & - & C & - & H \\ & & | & & | & & \\ & & H & & H & & \end{array}$$

In methane there are four C — H sigma bonds whereas in ethane there are six C — H and one C — C sigma bond.

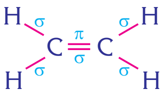
Formation of The Pi (π) Bond in The Ethylene Molecule

Both carbon atoms in ethylene molecule undergo sp^2 hybridization and form three identical sp^2 hybrid orbitals. One p orbital remains unhybridized. Two sp^2 hybrid orbitals from each carbon atom overlap end to end with the 1s orbital of a hydrogen atom and four C—H σ bonds are formed in total. Also, between the two carbon atoms, a C—C σ bond is formed as a result of the overlap between two sp^2 hybrid orbitals. So, in the C_2H_4 molecule in total there are five σ bonds. Meanwhile, the unhybridized p orbitals of the two carbon atoms overlap side by side and form a π bond. So between the two carbon atoms in the C_2H_4 molecule there is one σ bond, formed by the overlapping of sp^2 hybrid orbitals and one π bond, formed by the side by side overlapping of the unhybridized p orbitals. In total, two bonds are formed, hence a double bond exists between the two carbon atoms.

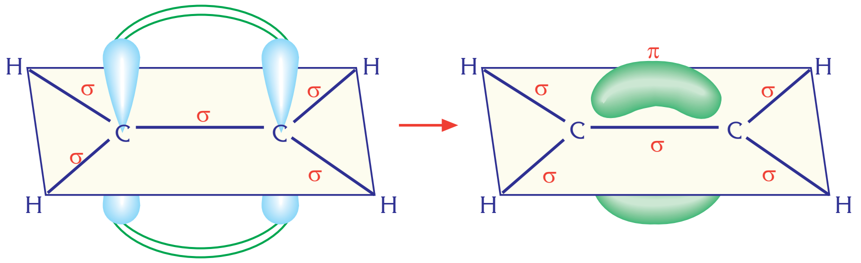
As a result, in the C_2H_4 molecule there are five σ and one π bond, so in total, six bonds.



When a carbon atom undergoes sp^2 hybridization, sp^2 hybrid orbitals form σ bonds, but the unhybridized p orbital forms a pi bond.



The ethylene molecule contains one C—C σ bond, four C—H σ bond and one C—C π bond.

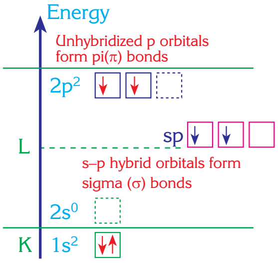


In the C_2H_4 molecule, unhybridized p orbitals overlap in side by side and form a π bond.

Formation of Pi (π) Bonds in The Acetylene Molecule

Both carbon atoms in the acetylene molecule undergo sp hybridization. Two p orbitals remain unhybridized. So, one sp hybrid orbital from each carbon atom overlaps with the s orbital of a hydrogen atom and two C—H σ bonds result. Also, between the two adjacent C atoms a C—C σ bond is formed as a result of end to end overlap of the sp hybrid orbitals. So in the C_2H_2 molecule there are three σ bonds in total.

Meanwhile, the unhybridized p orbitals of two carbon atoms overlap side by side and form two C—C π bonds. Thus, in the C_2H_2 molecule between the two carbon atoms, one σ bond is formed (by the end to end overlap of sp hybrid orbitals) and two π bonds are formed (by the side by side overlap of the unhybridized p orbitals).

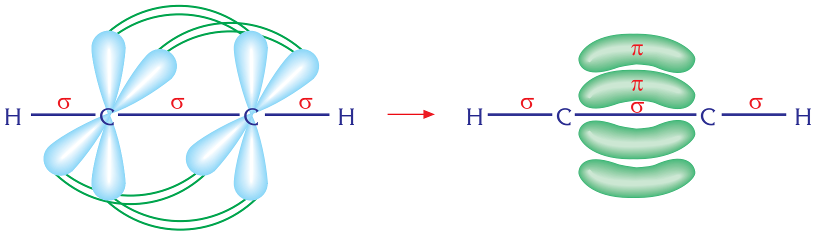


When carbon undergoes sp hybridization the sp hybrid orbitals form s bonds, and the unhybridized p orbitals form π bonds.

$H-C\equiv C-H$

Acetylene contains two C—H σ bonds, one C—C σ bond and two C—C π bonds. So in total there are three σ bonds and two π bonds.

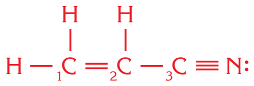
As a result, in the C_2H_2 molecule there are three σ and two π bonds, so in total five bonds.



In the C_2H_2 molecule, unhybridized p orbitals form two π bonds by overlapping side by side.

Example **17**

The formula of acrylonitrile which is a basic material in the production of synthetic fabrics is given below:



$$H-\overset{\displaystyle H}{\overset{|}{{}_1C}}=\overset{\displaystyle H}{\overset{|}{{}_2C}}-{}_3C\equiv N:$$

How many σ and π bonds are there in this molecule?

Solution

There are four single bonds (three C — H and one C — C) in the molecule. Those bonds are σ bonds. But, the molecule contains six σ bonds in total because both carbon – carbon double and carbon nitrogen triple bonds contain one σ bond.

The molecule contains three π bonds: one in the carbon–carbon double bond and two in the carbon nitrogen triple bond.

6. RESONANCE STRUCTURES

In some molecules there may be a conflict between the theoretical and real structures. For example, the structure of the ozone molecule (O_3) should contain one single bond and one double bond between the oxygen atoms according to our rules. Only in this case, each oxygen atom in the ozone molecule can complete its octet and obtain the configuration of a noble gas. So the structure of the ozone molecule should be.

EXTRA

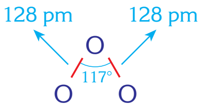
Bond Length

The distance between the nuclei of two bonding atoms in a molecule is called the bond length. The most important factor controlling the bond length is the radii of the atoms that form the bond. As the atomic radii of the atoms increase, the bond length will also increase.

Also, each π bond added to a σ bond makes the bond shorter. So the bond length between two carbon atoms decreases as π bonds are added.

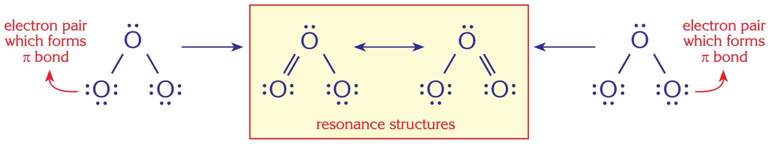
For bonds between any two atoms, increasing number of bonds decreases the bond length. As the number of bonding electron pairs increases, the attractive force between the atoms gets stronger. Therefore,

C — C 154 pm	C = C 134 pm	C ≡ C 121 pm
C — O 143 pm	C = O 122 pm	C ≡ O 113 pm

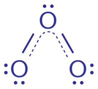


The bonds in the ozone, molecule, O_3, are identical and have a length of 128 pm.

A $O=O$ bond is shorter than a $O-O$ bond, but studies show that in the O_3 molecule both oxygen-oxygen bonds are of equal length. Moreover, this bond length is found to be shorter than a single bond but longer than a double bond. Therefore the structure of the molecule is a hybrid of the two molecular structures shown below.



A structure midway between the two resonance structures represents the ozone structure best. The bonds in this structure are stronger than a single bond but weaker than a double one.



Octet Rule

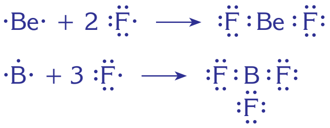
Atoms tend to acquire a noble gas configuration either by forming ions or by sharing electrons in covalent bonds. The tendency of atoms to acquire eight valence electrons is known as the octet rule.

7. EXCEPTIONS TO THE OCTET RULE

Atoms form bonds to make their electronic structures similar to those of noble gases. All noble gases except He have an electron structure ending with ns^2 np^6. Most atoms complete their valence shell with eight electrons (an octet) to become stable. However, some exceptions occur.

7.1. ELECTRON DEFICIENCY

Some atoms are able to form compounds even though the resulting structure doesn't provide eight valence electrons. For example beryllium and boron do not complete their octet in their covalent compounds because these atoms have less than four valence electrons. For example, in BeF_2; (F – Be – F) beryllium shares its two valance electrons but it doesn't complete its octet, it is only surrounded by four electrons. In BF_3, the boron atom shares its three valence electrons but does not complete its octet as it has just three electron pairs (six electrons) surrounding it.



The same principle applies for $BeCl_2$, BeH_2, BCl_3 etc. Beryllium and boron compounds are exceptions to the octet rule.

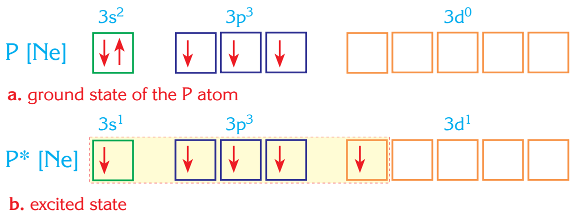
7.2. EXPANDED OCTETS

Some atoms in 3rd period of the periodic table and beyond may complete their octet and form a stable compound. They may also disobey the octet rule by having more than eight electrons in their valence orbitals.

There might be five or six electron pairs around an atom.

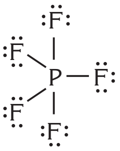
Expanded octet in the PF_5 molecule

The electron structure of phosphorus ends with $3s^2$ $3p^3$ $3d^0$. As we saw previously, the phosphorus atom normally forms three bonds. However, it is able to undergo sp^3d hybridization as shown below.

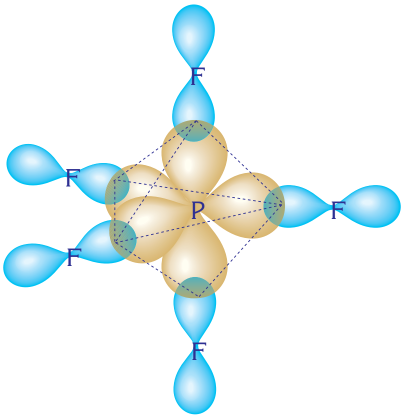


When it is excited one of the 3s electrons is promoted to a 3d orbital. In this configuration, phosphorous has five half–filled orbitals, and therefore a bonding capacity of five. When these half-filled orbitals are filled with the unpaired electrons from five fluorine atoms, the PF_5 molecule results. In this molecule, the phosphorous atom is surrounded by five pairs, or ten electrons.

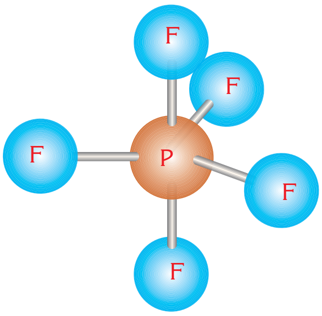
The Lewis structure of the PF_5 molecule is;



So PF_5 is an exception to the octet rule.



Orbital orientation in PF_5, the molecule is trigonal bipyramidal

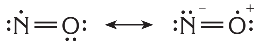


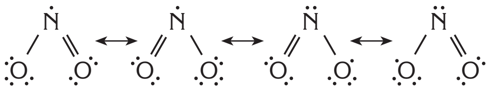
Molecular model of PF_5

Trigonal bipyramid

7.3. FREE RADICALS

Compounds that have unpaired electrons in their structures are called free radicals. These compounds also do not obey the octet rule.





NO and NO_2 are two examples of free radicals.

Free radicals are chemically active substances. They do not have any charge.

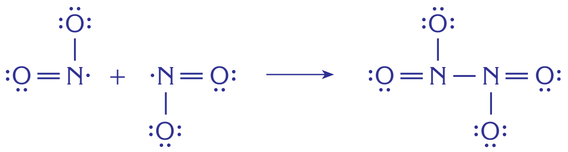
Example 18

Two NO_2 molecules may easily combine and form the N_2O_4 molecule. Explain the reason for this combinaton.

Solution

$2NO_2 \rightarrow N_2O_4$

The nitrogen atom in the NO_2 molecule has an incomplete octet, having a single unpaired electron. The unpaired electrons of the nitrogen atoms combine to form a single bond.



By the combination of two NO_2 molecules, the nitrogen atoms complete their octet and become more stable.

SUPPLEMENTARY QUESTIONS

1. Compare the electronegativities of the following elements; P, N, F, Si.

2. Why do atoms tend to acquire a noble gas electron configuration?

3. Explain the bond formation in Cl_2 and O_2 using orbital representation. ($_{17}Cl$, $_8O$)

4. Write the electron configurations of following species and draw their orbital diagrams.

a. $_7N^{-3}$ **b.** $_{12}Mg^{+2}$ **c.** $_{16}S^{-2}$ **d.** $_{26}Fe^{+3}$ **e.** $_{47}Ag^{+1}$

5. Draw the electron dot structures of the following elements

a. $_5B$ **b.** $_{12}Mg$ **c.** $_{15}P$ **d.** $_{19}K$ **e.** $_{35}Br$

6. Give the electron dot representations of the ions and compounds given below.

a. HF **b.** CO_2 **c.** C_2H_2 **d.** H_2S **e.** NCl_3

f. Cl^- **g.** CN^- **h.** ClO_2^- **i.** SO_3^{2-} **j.** PO_4^{3-}

($_1H$, $_7N$, $_8O$, $_9F$, $_{12}C$, $_{15}P$, $_{16}S$, $_{17}Cl$)

7. For the H_2S molecule, show its

a. Orbital representation.
b. Electron dot representation.
c. Line representation.

8. Which of the following compounds exhibit ionic bonding

a. H_2O **b.** Na_2O **c.** KCl **d.** $CaBr_2$ **e.** P_2O_5

9. Compare the ionic character of the given compounds.

a. $NaCl$ **b.** KF **c.** MgO **d.** CaS **e.** AlF_3

10. Show the formation of ionic bonds between

a. K and Cl **b.** Mg and F **c.** Be and O

11. What is the difference between the formation of ionic and covalent bonds?

12. Draw the molecular structures of the following species. Are these molecules polar or non-polar?

a. NH_3 **b.** CH_4 **c.** H_2O **d.** HF

e. PH_3 **f.** CO_2 **g.** PCl_3 **h.** BeH_2

13. Describe the type of bonds in each of the following compounds.

a. $AlCl_3$ **b.** SF_6 **c.** CCl_4 **d.** $NaNO_3$ **e.** $CaSO_4$

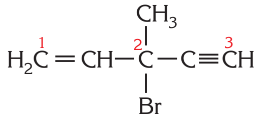
14. Explain coordinate covalent bonding and give one example.

15. Explain the bonding in the H_3O^+ and BF_4^- ions.

16. Show the coordinate covalent bond formed between BF_3 and NH_3 molecules?

The shape of the BF_3 molecule is trigonal planar but NH_3 molecule is trigonal pyramidal. Explain the reason for this difference.

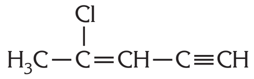
17. Write down the types of hybridization of the numbered carbon atoms in the following compound.



$$H_2\overset{1}{C}=CH-\overset{CH_3}{\underset{Br}{\overset{2}{C}}}-C\equiv\overset{3}{C}H$$

18. Explain the hybridization undergone by the boron atom when it bonds with hydrogen.

19.



$$H_3C-\overset{Cl}{C}=CH-C\equiv CH$$

What are the number of σ and π bonds in the above compound?

20. Compare the carbon - carbon bond lengths of given compounds

a. C_2H_6 **b.** C_2H_2 **c.** C_2H_4

21. The angle between the N – H bonds in the NH_3 molecule is 107°, whereas the angle between the H – O bonds in the H_2O molecule is 104,5°. What is the reason for this difference?

22. Find the number of π and σ bonds in each of the following molecules.

a. O_2 **b.** CO_2 **c.** N_2 **d.** C_2H_4 **e.** C_2H_2

MULTIPLE CHOICE QUESTIONS

Chemical Bonding and Molecular Structures

1. X^{+n} and Y^{-n} have the same number of electrons and a stable noble gas electron configuration. According to this information;

I. The atomic number of Y is 8 if the atomic number of X is 12.

II. Both X and Y elements are in the same period.

III. Both X and Y elements are in the same group.

Which of the above statements is/are correct?

A) I Only B) I and II C) I and III D) III Only E) I , II and III

2. Nitrogen and hydrogen molecules react to form ammonia. Which of the following statements is/are correct for the ammonia molecule?

I. The molecule is polar.

II. The N–H bonds in the molecule are polar.

III. The molecular structure is trigonal pyramidal.

A) I Only B) II Only C) I and II D) II and III E) I, II and III

3. I. H_2O

II. NH_3

III. CH_4

Which of the above molecules is/are polar?

A) I Only B) II Only C) I and II D) II and III E) I, II and III

4. Which one of the following molecules has a nonpolar bond ?

A) NaCl B) $MgCl_2$ C) $AlCl_3$ D) Cl_2 E) HCl

5. How many σ bonds does the C_3H_8 molecule have?

A) 6 B) 7 C) 8 D) 9 E) 10

6. Some molecules and their Lewis structures are given below ;

	Molecule	Lewis structures
I.	H_2O	$\mathrm{H}:\underset{\cdot\cdot}{\overset{\cdot\cdot}{\mathrm{O}}}:\mathrm{H}$
II.	NH_3	$\mathrm{H}:\underset{\mathrm{H}}{\underset{\cdot\cdot}{\overset{\cdot\cdot}{\mathrm{N}}}}:\mathrm{H}$
III.	CO_2	$:\underset{\cdot\cdot}{\overset{\cdot\cdot}{\mathrm{O}}}:\mathrm{C}:\underset{\cdot\cdot}{\overset{\cdot\cdot}{\mathrm{O}}}:$

Which of the given matching pairs above is/are correct?

A) I Only B) II Only C) I and II D) I, II and III E) II and III

7. Which shaped molecular structure does hydrogen form with a group 5A element?

A) Linear B) Angular C) Trigonal planar D) Tetrahedral E) Trigonal pyramidal

8. In which one of the following molecules is the bond between the C atoms the shortest ?

A) CH_4 B) C_2H_4 C) C_2H_2 D) C_3H_8 E) C_4H_{10}

9. I. It contains two σ bonds

II. It contains two π bonds

III. It is bent

Which of the above statements is/are correct for a molecule of CO_2?

A) I Only B) II Only C) III Only D) I and II E) I and III

10. In which one of the following molecules does the central atom undergo sp^2 hybridization?

A) $MgCl_2$ B) BF_3 C) NaCl D) CH_4 E) CO_2

11. Which of the following molecules is tetrahedral?

A) BH_3 B) H_2S C) CCl_4 D) OF_2 E) NH_3

12. Which explanation is incorrect regarding the group 5A elements.

I. Compounds of phosphorus cannot form five bonds.

II. When it is excited, the phosphorus atom becomes pentavalent.

III. Nitrogen can form both ionic and covalent bonds.

A) I only B) II only C) III only D) I and II E) I and III

13. The respective bond types in Cl_2, HCl and NaCl are:

A) Ionic, polar covalent, non-polar cavalent

B) Non-polar covalent, polar covalent, ionic

C) Polar covalent, non-polar covalent, ionic

D) Ionic, non-polar covalent, polar covalent

E) Non-polar covalent, ionic, polar covalent

14. What is the oxidation state number of carbon in CO_2?

A) 1 B) 2 C) 3 D) 4 E) 5

15. How many electrons are used in the bond formation of the hydrogen molecule?

A) 1 B) 2 C) 4 D) 6 E) 8

16. How many polar and nonpolar covalent bonds are there between the atoms in the ethene molecule C_2H_4?

A) 2 polar, 2 nonpolar

B) 1 polar, 5 nonpolar

C) 4 polar, 2 nonpolar

D) 2 nonpolar, 4 polar

E) 5 polar, 1 polar

17. In which one of the following species would you expect to find a π bond?

A) H_2 B) O_2 C) F_2 D) Cl_2 E) Br_2

18. Which one of the following statements is wrong?

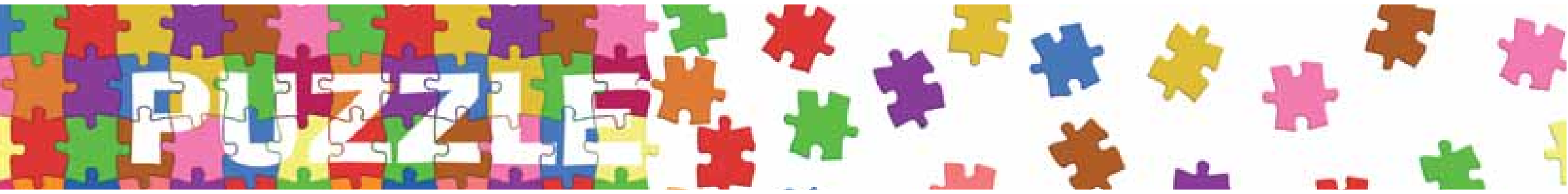
A) Chlorine is able to form both ionic and covalent bonds.

B) 2 electrons are used in the formation of the chemical bond in the chlorine molecule.

C) The bond in KCl is formed by electron exchange.

D) The bond types in HCl and NaCl are the same.

E) The bond type in Cl_2 is non-polar covalent.

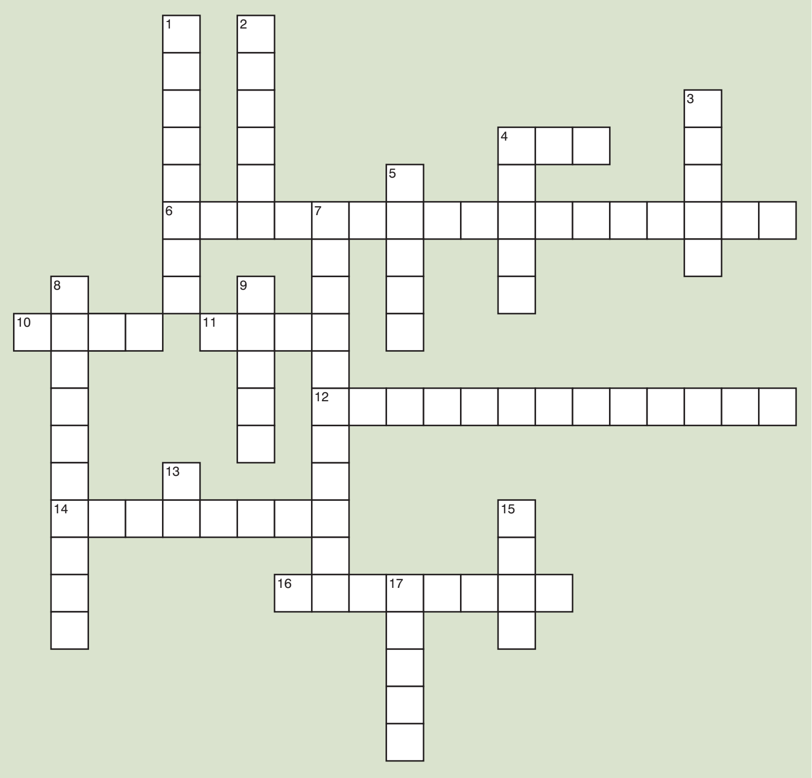


CRISS – CROSS PUZZLE

Complete the crossword in the normal way.

CLUES ACROSS

4. This is the type of hybridization of the carbon atom in the methane molecule.
6. This is the tendency of atoms to attract bonding electrons within a molecule.
10. This is the attractive force that holds atoms together in a compound.
11. The number of electrons necessary to form a double bond.
12. This is the combination of pure atomic orbitals to produce new orbitals.
14. These type of molecules do not have oppositely charged poles.
16. This is the most electronegative element.



CLUES DOWN

1. When two non-metals share their electrons, the result is this bond.
2. Such a bond exists between the atoms of the nitrogen molecule.
3. The American scientist who proposed the theory of electron sharing.
4. Such bonds are formed by the end to end overlap of two orbitals.
5. This is a type of bond seen between two non-metals having different electronegativities.
7. A methane molecule has this shape.
8. In this type of covalent bond both of the bonding electrons are supplied by one atom.
9. A bond often seen between a metal and a non-metal.
13. Such orbitals have 50% s and 50% p character.
15. It is the shape of the water molecule
17. According to this rule, an atom tends to lose or gain electrons until it has eight electrons in its valence shell.

CRYPTOGRAM

Below is a phrase about bonding. Try to find out the whole phrase with the given clues.

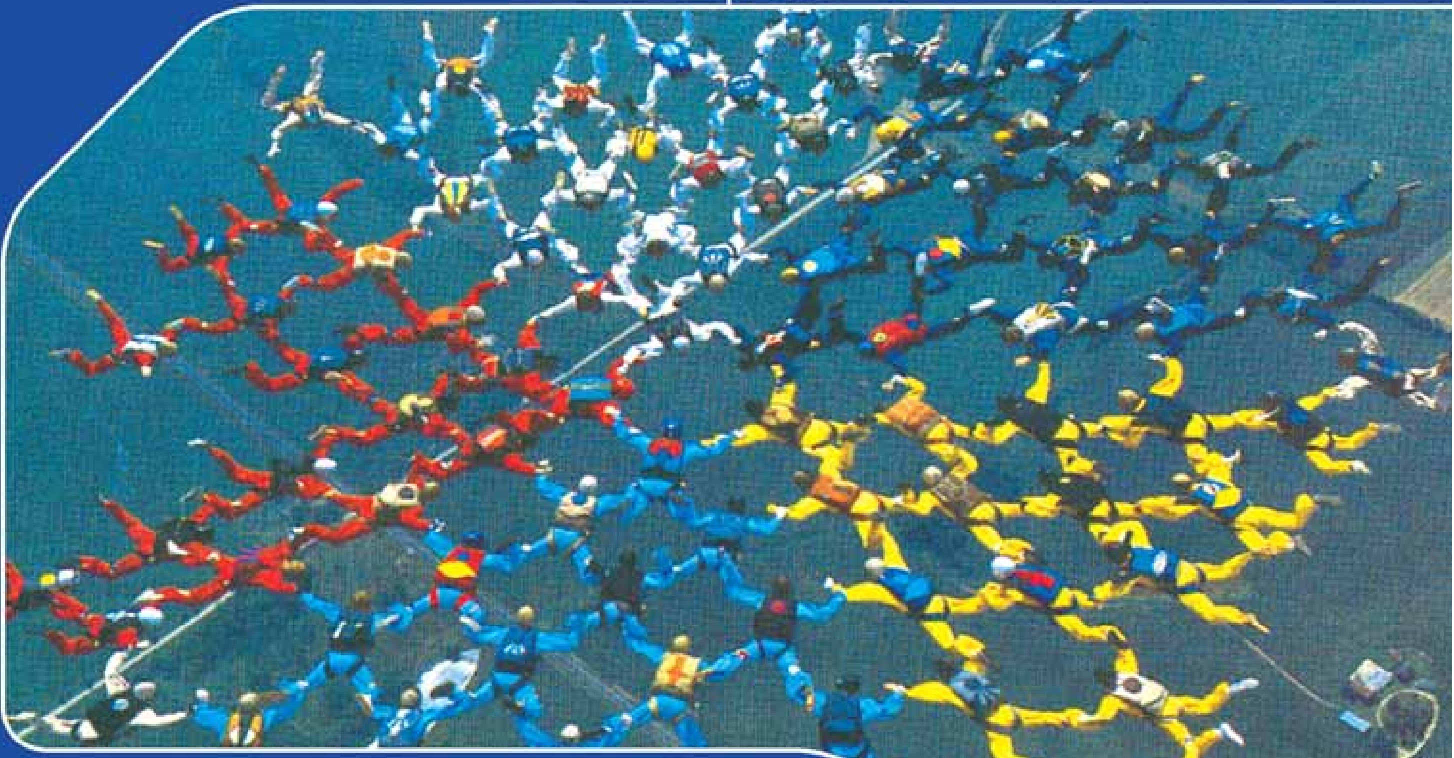
A	B	C	D	E	F	G	H	I	J	K	L	M	N	O	P	Q	R	S	T	U	V	W	X	Y	Z
		3			13		22	4						23					19						

3 22 2 8 4 3 18 21   20 23 10 7   4 24   19 22 2   18 19 19 14 18 3 19 4 23 10 18 21

13 23 14 3 2   19 22 18 19   22 23 21 7 24   18 19 23 8 24

19 23 12 2 19 22 2 14   4 10   3 23 8 16 23 11 10 7 24

BONDS IN SOLIDS & LIQUIDS



CHEMICAL BONDS

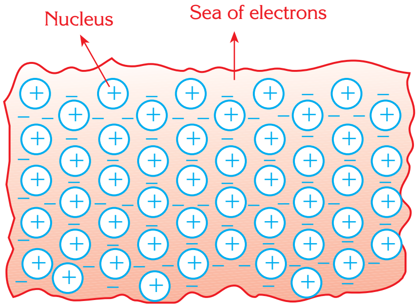
In the gaseous state molecules are far apart from each other and move continuously. When the temperature is decreased, the molecules slow down and lose kinetic energy, as a result the molecules can stick together and the physical state changes to liquid or solid.

INTRODUCTION

All gases condense at low temperatures to become liquids. If the temperature is lowered still further, liquids turn into solids.

In the solid and liquid phases, molecules are very close to each other. This is because forces hold the molecules together in the solid and liquid states. We have already studied intramolecular bonds within molecules in the previous chapter. In this chapter, we will examine the forces of attraction between the particles in liquids and solids.

These forces affect the boiling point, melting point, hardness, and electrical and heat conductivity of a substance. In this chapter, we will study metals, ionic solids, network solids, dipole-dipole attractions, van der Waals forces and hydrogen bonds.



In metals, an attraction is formed between the negatively charged electrons and the positively charged nuclei of the atoms.

1. METALLIC BONDS

Metal atoms have a small number of valence electrons. The nuclear attractive forces between the metal nuclei and their valence electrons are reduced by the inner electrons (which are closer to nucleus). Thus, the nucleus of a metal atom exerts only a small attractive force on its valence electrons and these electrons are able to move more freely. For this reason, metal atoms have very low ionization energies and electronegativities.

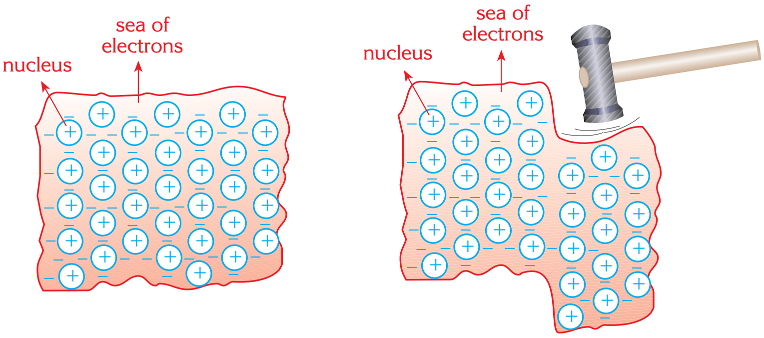
Metals are solid at room temperature, except for mercury. This tells us that the attractive forces between metal atoms are strong. The valence electrons of metal atoms can easily move from the free orbitals of one atom to another. These electrons that can move freely between atoms form an "electron sea". An attractive force occurs between the negatively charged "sea of electrons" and the positively charged nuclei. Metal atoms are held together because of this attractive force. This is called the ***metallic bond***.

Metals can be drawn into wires and hammered easily.

Let's examine sodium. The electron configuration of sodium is;

The Na atom has one half-filled ($3s^1$) and three empty orbitals ($3p_x$, $3p_y$, $3p_z$). The number of valence orbitals is greater than the number of valence electrons. In the solid state, sodium atoms are surrounded by other sodium atoms. Thus, the valence electron of the sodium atom in the 3s orbital can move to the empty orbitals ($3p_x$, $3p_y$, $3p_z$) of neighbouring atoms. When each sodium valence electron behaves in this way, a sea of electrons is built up around the sodium atoms (now positive ions, having lost a valence electron).

Because of the attraction between the electron sea and the positively charged sodium nuclei, a metallic bond is formed. Because of these freely moving electrons in the electron sea, metals are good conductors of heat and electricity. They can be drawn into wires and can be hammered into shape easily.



The free movement of electrons in metals makes it easy for metals to be shaped and drawn into wires.

In the periodic table, metallic bond strength generally decreases as you go down a group.

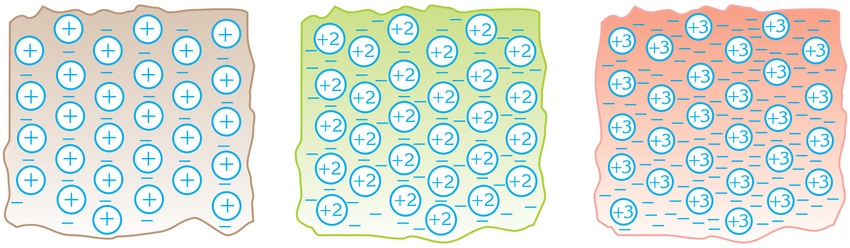
However across a period, the metallic bond strength generally increases from left to right. This is because the metals on the right hand side possess a higher number of valence electrons.

Let us compare the metallic bonds of sodium, magnesium and aluminum. sodium ($_{11}$Na) has one valence electron, magnesium ($_{12}$Mg) has two and aluminum ($_{13}$Al) three valence electrons.

The metallic bonding in Al is the strongest, and the weakest in Na.

Strong metallic bonds increase the boiling point, melting point, and the hardness of the metal.

Sodium, magnesium and aluminum melt at 98°C, 650°C and 660°C respectively.



sodium magnesium aluminum

Comparison of metallic bonding in Na, Mg and Al.

Sodium

Magnesium

Aluminum

Example 1

The melting points of sodium and calcium are 98°C and 838°C respectively. Give the reason for this.

Solution

Na and Ca both contain metallic bonds. The strength of a metallic bond is related to the number of valence electrons.

Na has one valence electron and Ca has two valence electrons therefore the total charge of the electron sea in Ca is greater than that of Na. So the metallic Ca bonds are stronger than those in Na, and therefore Ca melts at a higher temperature.

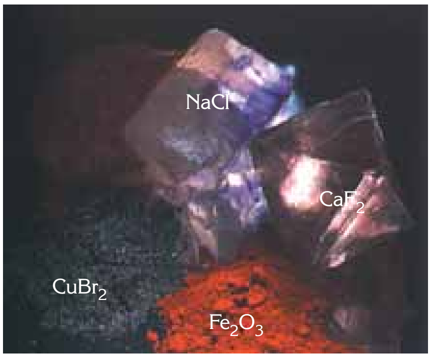
2. IONIC SOLIDS

Some ionic crystals. Nickel(II) nitrate ($Ni(NO_3)_2 . 6H_2O$, green), potassium dichromate ($K_2Cr_2O_7$, orange), copper(II) sulfate ($CuSO_4 . 5H_2O$ blue)

When metal and nonmetal atoms come together they form ionic bonds, as you will remember from the previous chapter. In the ionic bond, the metal atoms which lose electrons become positively charged and the nonmetal atoms, which gain electrons, become negatively charged. Electrostatic attraction occurs between the positive and negative charges, holding the ions together.

These electrostatic attractions act in all directions. Thus, ionic crystalline solids consist of metal ions are surrounded by non-metal ions and non-metal ions surrounded by metal ions. Therefore, ionic solids do not have a molecular structure.

As the attraction between the ions is strong, the melting and boiling points of ionic solids are very high. For example, NaCl melts at 801°C.



Some salts crystals.

In ionic solids, electrons are held in place around the ions so they don't conduct electricity. However, in aqueous solution and molten state, they do conduct electricity. Electrical conductance of ionic compounds is not due to movement of electrons but to the movement of ions.

Ionic compounds are brittle but not ductile, as is shown in Figure 1. When they are hammered, their structure is disturbed, the hammered part shifts and similar charged ions repel each other and the ionic substance breaks down into smaller pieces. Since movement of the ions disturbs the balance of electrical charge, ionic solids cannot be drawn into wires and are broken easily.

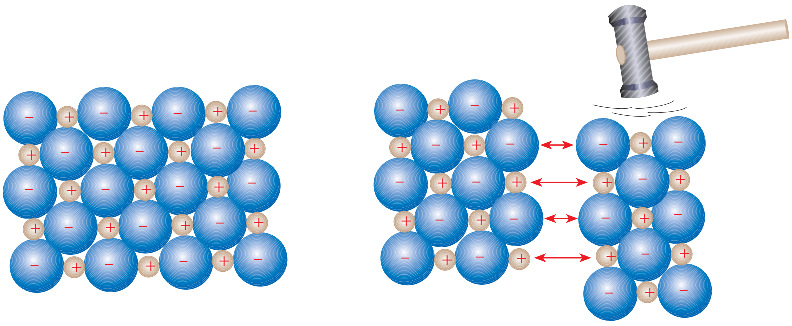


Figure 1 : *Ionic solids are brittle and can not be drawn into wires or hammered into plates. As the ionic bonds are very strong, compounds containing such bonds are very stable.*

Example **2**

Compare the ionic character of the following salts. KF, KCl, KBr and KI

Solution

Electronegativity decreases from top to bottom down a group. The electronegativity order for the halogens is F > Cl > Br > I.

Among the given compounds KF is the most ionic and KI is the least ionic.

3. NETWORK SOLIDS

In molecular covalent compounds, intermolecular forces are very weak in comparison with intramolecular forces. For this reason, most covalent substances with a low molecular mass are gaseous at room temperature. Others, with higher molecular masses may be liquids or solids, though with relatively low melting and boiling points.

However, in some covalent substances, known as network solids, atoms are bonded together in a way that forms a network structure.

Diamond

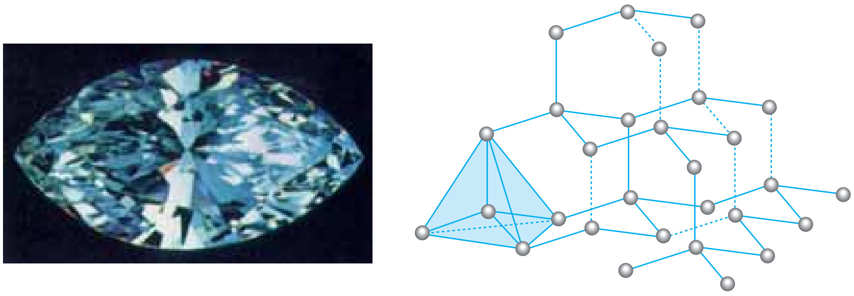
The most typical example of a network solid is diamond. In diamond each carbon atom is covalently bonded to four other carbon atoms forming a tetrahedral shape. (The type of hybridization that corresponds to this tetrahedral structure is sp^3) This structure is extremely strong and this makes diamond the hardest natural substance.

Silicon carbide SiC is another network solid. Silicon carbide is used as an abrasive because of its hard structure.

EXTRA

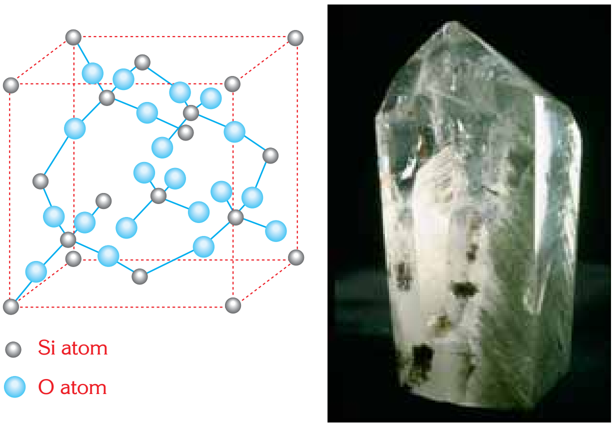
Allotropy

Diamond and graphite are allotropes of carbon. The density of diamond is 3.5 g/cm^3 and that of graphite is 2.2 g/cm^3. Diamond is used to cut other hard materials such as glass because of its hardness. On the other hand, softer graphite is used in pencils.

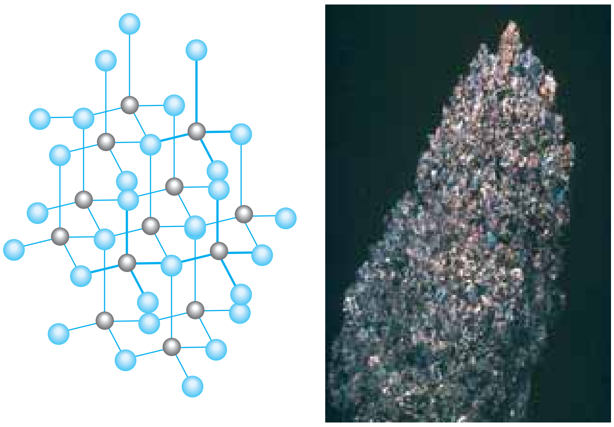


Each carbon atom is bonded to 4 other carbon atoms to form a tetrahedral shape in diamond. The bonds are formed by sp^3-sp^3 hybrid overlap.

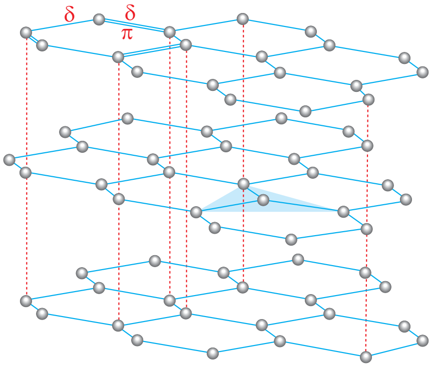
Diamond and silicon carbide are nonconductors of electricity and have very high melting points. The melting point of diamond is about 3500°C and that of SiC 2830°C.



Quartz (SiO_2) has a network covalent bond structure



The network covalent bond in silicon carbide (SiC)



Carbon atoms in graphite form a hexagonal structure.

Graphite

In graphite, a different form of carbon, atoms are bonded to each other in such a way that a hexagonal structure is formed in a plane. Each carbon atom is bonded to three other carbon atoms with an angle of 120° between the bonds. The bonding involves $sp^2 - sp^2$ hybrid overlap and this gives rise to layers.

Bonds in the same plane are very strong, but attractions between the layers are much weaker. Because of this weak bonding between the layers, the layers can slide over each other. This makes graphite a good lubricant, and gives it a soft feel. It conducts electricity.

EXTRA

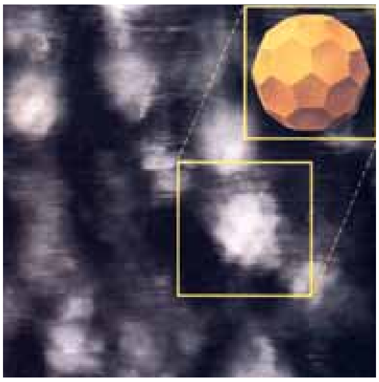
Fullerenes

In 1985, another allotropic form of carbon, fullerenes, were discovered. Two important fullerenes are C_{60} and C_{70}.

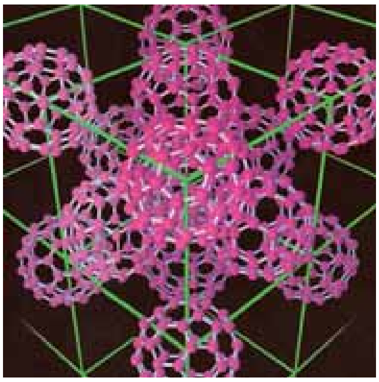
The C_{60} molecule is produced by evaporating carbon at 8000°C. The structure of C_{60} resembles a soccer ball and contains 12 pentagons and 20 hexagons. Carbon atoms are located at the corners of these interlocking shapes.

Today fullerenes are used in production of batteries, treatment of cancer and in electronic devices.

The C atoms in C_{60} are arranged as in a football.



A view of C_{60} molecules on silicon crystals (the white patches) obtained from a scanning tunnelling microscope.



Model of the C_{60} crystal drawn from the view obtained from the scanning tunnelling microscope.

Example 3

Explain whether or not the following substances conduct electricity.

a. Al (s), b. KCl(aq), c. Hg(l), d. $NaNO_3$(s), e. SiC(s), f. LiF(aq)

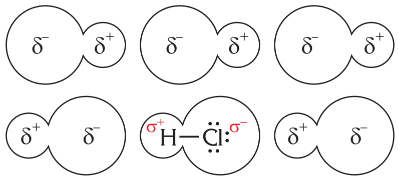
Solution

Al(s) and Hg(s) are metals and consist of metal ions in an electron sea. Although their physical phases are different, they conduct electricity because of the freely moving electrons that make up the electron sea.

KCl, $NaNO_3$ and LiF are ionic solids. Ionic solids don't conduct electricity in the solid state, however, aqueous solutions and molten forms of ionic compounds contain mobile ions so they can conduct electricity. Thus, aqueous KCl and LiF conduct electricity but solid $NaNO_3$ doesn't.

SiC(s) is a network covalent solid. It contains covalent bonds between its atoms. It doesn't have any freely moving electrons or ions and so SiC doesn't conduct electricity.

4. DIPOLE – DIPOLE FORCES



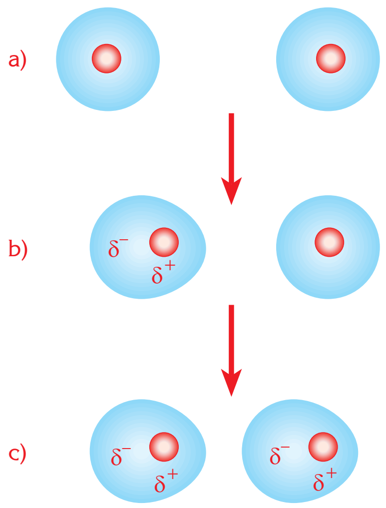
Formation of dipole – dipole forces..

In polar covalent substances, the molecules have partial positive and negative charges because of the electronegativity differences between the atoms. The molecules are said to possess a dipole.

There is an attraction between the positive end of one dipole and the negative ends of neighboring dipoles. This attraction is called **dipole – dipole** attraction.

For example, in the HCl molecule, the partial charge on the hydrogen atom is positive and the partial charge on the chlorine atom is negative. Between neighboring HCl molecules there is an attraction between the hydrogen and chlorine ends of the molecules. Dipole - dipole forces between HCl molecules are much weaker than the covalent bond within the molecule.

5. VAN DER WAALS FORCES



Formation of van der Waals forces. Molecules approaching each other gain momentarily polar character.

Noble gases and non-polar molecules such as CO_2 and CH_4 do not have dipoles. In these molecules, the movement of electrons results in nonpolar molecules becoming temporarily polar; an instantaneous dipole is formed. The molecule which becomes momentarily polar then causes its neighboring molecule to become polar. Thus a weak attraction occurs between the molecules. This attraction is named the van der Waals force.

Van der Waals forces depend upon the electron density of the atoms. Increasing number of atoms in a molecule increases the van der Waals attractive force. Since the electron number of a neutral atom is equal to its proton number, atoms which have a large proton number have strong van der Waals forces between their molecules. Therefore, **van der Waals** forces are stronger between molecules with high molecular masses.

Van der Waals forces between I_2 molecules are stronger than those between Cl_2 molecules because clearly, iodine has bigger molecules than chlorine. Propane (C_3H_8) is bigger than methane (CH_4), so the van der Waals forces between C_3H_8 molecules are stronger than those between CH_4 molecules.

For small molecules, the van der Waals force is weaker than dipole - dipole forces and hydrogen bonding. Thus, small nonpolar molecules have low melting and boiling points.

Let's compare the intermolecular forces between I_2 and Cl_2 . I_2 has the greater molecular mass so the van der Waals forces between its molecules are greater in comparison with Cl_2. Therefore at room temperature iodine is solid whereas chlorine is gas.

The first container contains bromine and the second iodine.

Example 4

Compare the boiling points of the following compounds:
CBr_4, CH_4, CF_4, CCl_4.

Solution

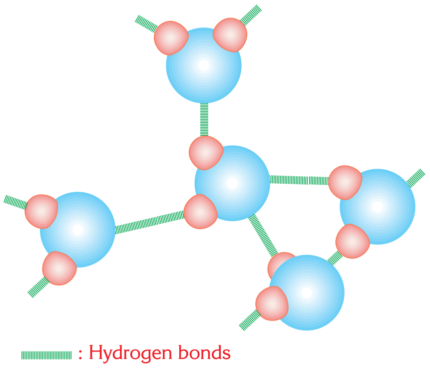
Boiling points of substances increase with increasing intermolecular forces. All the given compounds are non-polar. We know that the non-polar molecules possess van der Waals forces and these forces are proportional to the molecular masses of the compounds. Therefore CH_4, having the smallest molecular mass, has the lowest boiling point. So the boiling point order is;

$CH_4 < CF_4 < CCl_4 < CBr_4$

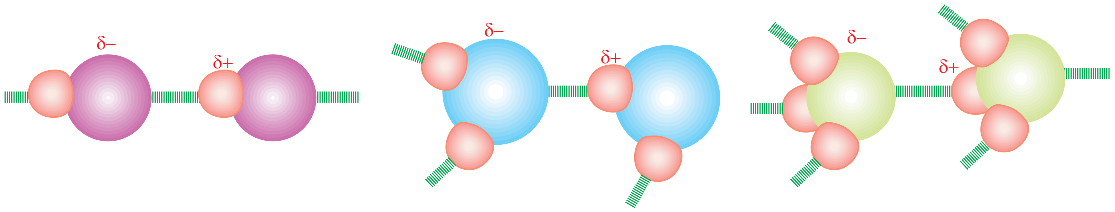
6. HYDROGEN BONDS

Fluorine, oxygen and nitrogen are the most electronegative elements. Therefore the compounds that these elements form with hydrogen (HF, H_2O,NH_3) are highly polar. Due to this polarity an intermolecular force that is much stronger than the usual dipole-dipole attraction occurs. These strong intermolecular forces are called **hydrogen bonds**.

A hydrogen bond is formed between a hydrogen atom and a lone pair electrons from an atom in a neighboring molecule. For example, the hydrogen atom of a water molecule forms a hydrogen bond with the lone pair of electrons from an oxygen atom in another water molecule.



Hydrogen bonds between water molecules are stronger than dipole-dipole and van der Waals forces.

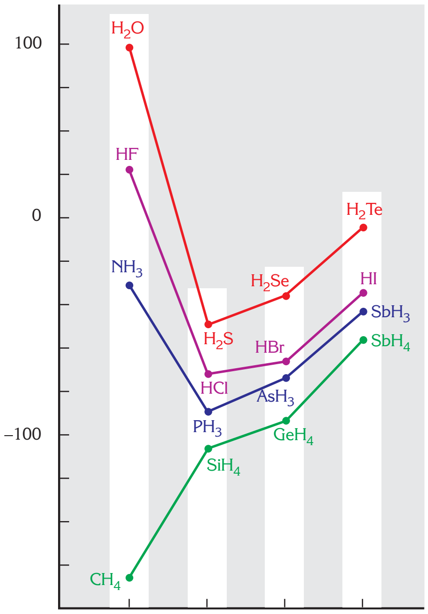


HF molecules

H_2O molecules

NH_3 molecules

Comparing the hydrogen bonds of HF, H_2O and NH_3; HF > H_2O > NH_3

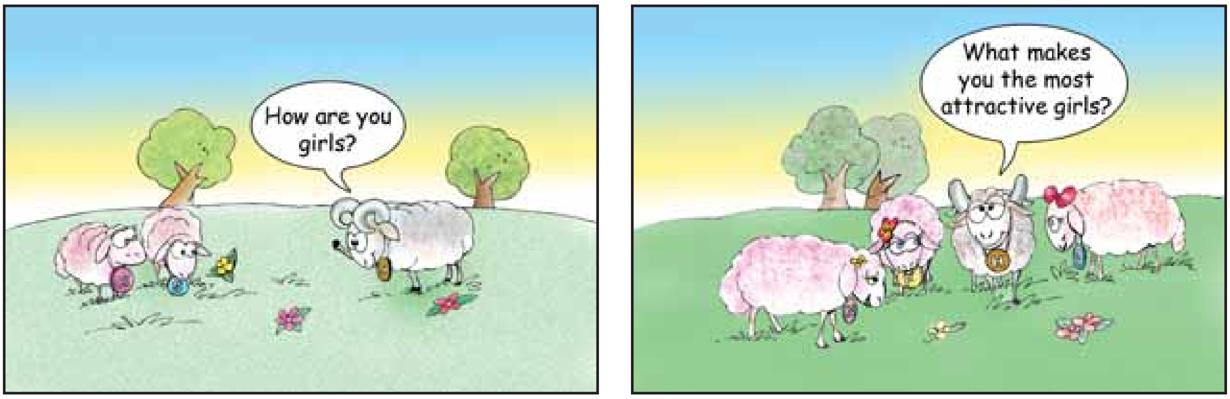


Because of hydrogen bonding the boiling points of HF, H_2O and NH_3 are greater than expected.

Although there are van der Waals forces between water molecules, the effect of the hydrogen bonding is much stronger than that of the van der Waals forces. For this reason, the boiling point of water is higher than expected.

If the boiling point of a substance is high, this tells us that the intermolecular forces in this substance are also high.

The boiling points of the hydrides of the group 4A elements (CH_4, SiH_4 and SnH_4) increase gradually with increasing atomic number. Other groups (VA, VIA and VIIA) show the same general trend, however, NH_3, H_2O and HF show an unexpected increase in boiling point. This is explained by the fact that these molecules have hydrogen bonding occurring between them. The boiling points don't show the expected pattern. For example, if the curve that takes in H_2Te, H_2Se, H_2S is extended to the second period; the boiling point of water would be expected to be around –90°C. However, the boiling point of water is 100°C and so it can be summarized that hydrogen bonding increases the boiling point of water by around 190°C.



Fluorine, oxygen and nitrogen are the most electronegative elements.

Example 5

What is the main type of bonding that must be overcome to carry out the changes of state given below?

a. Fe(s) → Fe(l), **b.** $H_2O(l) \rightarrow H_2O(g)$, **c.** $I_2(s) \rightarrow I_2(l)$

d. $AlCl_3(s) \rightarrow AlCl_3(l)$, **e.** $S_8(s) \rightarrow S_8(l)$

Solution

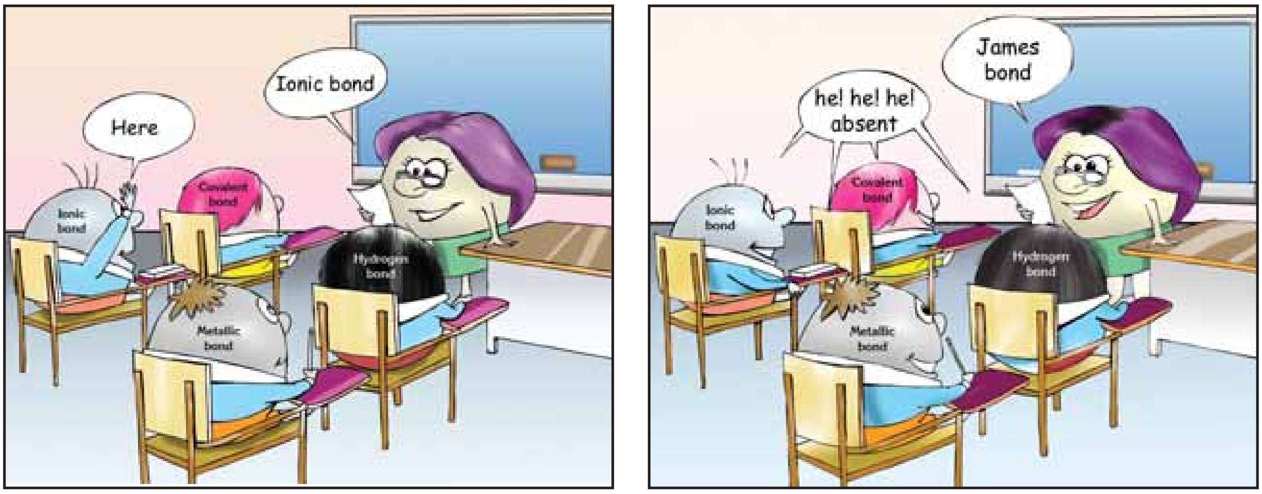
a. Since iron is a metal, metallic bonds must be overcome.

b. $H_2O(l)$ contains hydrogen bonds between its molecules. Thus hydrogen bonds must be overcome.

c. I_2(iodine) is a nonpolar molecule and has van der Waals forces between the molecules. So these bonds need to be overcome.

d. $AlCl_3(s)$ is an ionic crystal. The ionic bonds which form the ionic crystal must be overcome.

e. S_8, has van der Waals forces between its molecules, these bonds must be overcome.



READING

HOW DOES AN IRON WORK?

Fabric is a flexible, artificial substance made up of a network of natural or artificial fibres. It is formed by interlacing loops of yarn or thread and matting the fibers together by heat and pressure.

When these threads are loosened the flat form of fabrics is damaged.

Cotton fabrics are made up of cellulose molecules. These molecules are bonded to each other by hydrogen bonds which are easily broken by a sufficient amount of heat and water. These broken bonds (between the threads) can be bound together again using an iron. The iron breaks down the hydrogen bonds and then bonds them back together in a regular way.

Substance		Intermolecular Forces	Attracting Particles	Physical Properties
Network crystals (Ex : C, SiC, SiO_2)		Covalent	Atoms	– very high melting point – very hard – do not conduct electricity (except graphite)
Metals (Ex : Li, Cu, Pt, Fe, Hg)		Metallic	Positive cations and mobile electrons	– hard or soft – high melting point – malleable and ductile – conduct heat and electricity
Ionic Compounds (Ex : NaCl, $BaCl_2$, KNO_3)		Electrostatic	Positive and negative ions	– hard and brittle – high melting point – aqueous solutions and molten states conduct electricity
Molecular Compounds	Ex : H_2O, NH_3	– Van der Waals – Dipole - dipole – Hydrogen bond	Polar molecules. (partially negative and positive atoms)	– soft – low melting point – nonconductors or poor conductors of electricity
	Ex : H_2S, SO_2	– Van der Waals – Dipole - dipole		
	Ex : H_2, Cl_2, BF_3, CCl_4	Van der Waals	Nonpolar molecules. (momentarily partially negative and positive atoms)	

Table 1: *A summary of intermolecular forces*

SUPPLEMENTARY QUESTIONS

1. Explain why alkali and alkaline earth metals melt at lower temperatures than transition metals?

2. Arrange the alkali metals according to their metallic bond strength.

3. Arrange the given metals in order of increasing melting points.

I. $_{11}Na$
II. $_{12}Mg$
III. $_{19}K$

4. Which of the compounds in the following pairs has the greatest ionic character?

I. KF and KI
II. NaCl and KCl

5. Define network covalent bonds; list four substances that have a network covalent structure.

6. The boiling points of silicon carbide and silicon oxide are very high. Explain why this is so.

7. What are the differences in bonding between diamond and graphite?

8. Explain dipole - dipole attractive forces.

9. What are van der Waals forces? What are the factors that affect van der Waals forces?

10. Which forces are found in the following compounds?

I. LiCl
II. Na_2O
III. CCl_4

11. Compare the following substances according to their electrical conductivity? Give your reasoning.

a. sugar solution
b. table salt solution
c. vinegar

12. Define hydrogen bonds. What are differences between hydrogen bonds and dipole - dipole forces?

13. For each of the given pairs determine the substance with higher boiling point. Give your reasoning.

a. C_3H_8; C_5H_{12}
b. CH_4; CH_3OH
c. O_2; S_8
d. N_2; F_2
e. diamond, graphite
f. Cl_2; F_2
g. SiO_2; CO_2
h. He; Ar

14. Which one of the substances in the given pairs of compounds has the higher melting point?

a. Br_2, I_2
b. NH_3; NF_3
c. Ne; N_2
d. ICl; Cl_2
e. SCl_4; SCl_6
f. Ne; Ar
g. HCl, H_2
h. Fe; F_2
i. Na_2S; SO_2

15. Which forces must be overcome in order to melt the following species.

a. Zn
b. Ga
c. NaCl
d. $CaCl_2$
e. PF_3
f. NCl_3
g. NH_3
h. SO_2
i. MgO
j. F_2
k. Na
l. SiC

16. Which of the following substances have only van der Waals forces between their molecules?

a. Ne
b. CH_4
c. O_3
d. $MgCl_2$
e. CO_2
f. CO
g. Na_2O

MULTIPLE CHOICE QUESTIONS

1.

	Name	Bond Type
I.	Fluorine	Van der Waals
II.	Hydrogen fluoride	Hydrogen bond
III.	Lithium fluoride	Ionic bond

Above, a type of attractive force is given for each substance. Which of the matches is/are correct?

A) I only B) II only C) I and II D) II and III E) I, II and III

2. Which one of the following contains only van der Waals forces between its molecules?

A) CO_2 B) C_2H_5OH C) HCl D) H_2O E) KCl

3. Which of the following has the highest melting point?

A) Sodium B) Iodine C) HF D) Graphite E) Ice

4. The strength of intermolecular forces affects the boiling points of substances.

Which one of the following substances has the highest boiling point?

A) H_2 B) H_2O C) Na D) SiC E) CH_3OH

5. I. Be and H_2 have metallic bonds

II. H_2 has van der Waals forces between its molecules

III. Be and H may form compounds with dipole-dipole attractions between them.

Which of the above statements is/are correct?

A) I only B) II only C) I and II D) II and III E) I, II and III

6. I. Liquid helium

II. Liquid nitrogen

III. Liquid iron

Which of the above substances contain van der Waals forces?

A) I only B) II only C) I and II D) I and III E) I, II and III

7. Under room conditions

I. H_2 is a gas.

II. C_2H_5OH boils at a higher temperature than C_3H_8.

III. The boiling point of CH_4 is higher than that of H_2O.

Which of the above statements is/are correct?

A) I only B) II only C) I and II D) II and III E) I, II and III

8. Which of the following matches is **wrong**?

	Substance	Bond Type
A)	Neon	Van der Waals
B)	Bronze	Metallic
C)	Methyl alcohol	Ionic
D)	Table salt	Ionic
E)	Silicon carbide	Network covalent

9. Which of the following is/are correct for the bromine molecule?

I. It is a nonpolar molecule.

II. Van der Waals forces exist between the molecules.

III. Its intramolecular bonds are nonpolar covalent.

A) I only B) II only C) I and II D) II and III E) I, II and III

10. Which one of the given compounds is different from the others in terms of its electrical conductivity?

A) NaCl(l) B) Hg(l) C) KCl(s)

D) NaCl(aq) E) HCl(aq)

11. Which of the following species has the highest boiling point?

A) HCl B) CH_4 C) CO_2 D) HF E) Ne

12. Which of the following elements has a network covalent structure?

A) $_4Be$ B) $_8O$ C) $_{11}Na$ D) $_6C$ E) $_{17}Cl$

13. Which of the following gives the correct order for the melting points of $_{11}Na$, $_{13}Al$ and $_{18}Ar$?

A) Na>Al>Ar B) Al>Na>Ar C) Ar>Al>Na

D) Al>Ar>Na E) Na>Ar>Al

14. Which of the following gives the correct order for the boiling points of F_2, Cl_2 and Br_2.

A) $F_2>Cl_2>Br_2$ B) $F_2>Br_2>Cl_2$ C) $Cl_2>F_2>Br_2$

D) $Br_2>Cl_2>F_2$ E) $Br_2>F_2>Cl_2$

15. Which of the following compounds does not have hydrogen bonding between its molecules?

A) HF B) NH_3 C) H_2O D) C_2H_5OH E) HCl

16. Which one of the following compounds does not have dipole-dipole forces between its molecules?

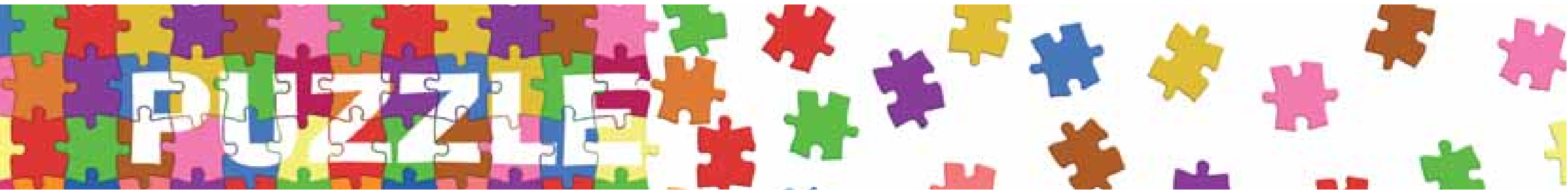
A) HCl B) $MgCl_2$ C) H_2S D) PH_3 E) NF_3

17. In which of the following are the intermolecular forces listed correctly from weakest to strongest for small molecules?

A) Van der Waals, hydrogen, dipole – dipole
B) Van der Waals, dipole – dipole, hydrogen
C) Hydrogen, Van der Waals, dipole – dipole
D) Hydrogen, dipole – dipole, Van der Waals
E) Dipole – dipole, Van der Waals, hydrogen

18. Hydrogen fluoride boils at 19.5°C while hydrogen chloride boils at –84°C. Which one of the following statements accounts for this difference?

A) HCl has greater van der Waals forces
B) HF has intermolecular hydrogen bonds.
C) HF has greater van der Waals forces
D) HCl has intermolecular hydrogen bonds
E) HCl has ionic bonds.



PUZZLE

WORD SEARCH

Try to find out the hidden words given below.

- BOND
- CRYSTAL
- DIPOLE
- GRAPHITE
- HYDROGEN
- IONIC
- METALLIC
- NETWORK
- OCTET

D	I	P	O	L	E	D	I	P	N	O	L	E	F	E
O	R	C	E	S	O	C	C	U	E	R	B	E	T	L
W	E	E	N	P	O	L	A	R	G	M	O	L	E	O
C	U	L	E	S	T	K	W	W	O	I	N	L	Q	P
C	R	Y	S	T	A	L	R	T	R	I	V	G	M	I
N	E	T	W	O	R	K	D	J	D	R	T	E	T	D
I	I	E	L	Q	W	Q	A	J	Y	N	T	I	J	Q
C	U	L	L	B	Y	A	T	D	H	A	O	U	E	O
K	U	L	U	O	D	J	P	E	L	T	M	B	T	S
C	N	V	C	A	X	W	J	L	Y	U	Q	E	I	Q
S	I	T	Z	S	K	I	I	O	H	I	A	F	H	E
D	E	N	E	I	M	C	O	D	B	S	V	U	P	B
T	Z	K	O	G	J	H	B	T	M	D	D	Z	A	V
I	Z	I	J	I	C	U	V	M	J	P	X	B	R	J
N	R	C	K	J	M	A	B	H	I	W	K	U	G	F

DOUBLE PUZZLE

The following words have been jumbled up, you must enter the correct spellings and tranfer the numbered letters to the secret message.

HDNYRGOE DOBN	11	☐	1	29	☐	☐	☐	☐		☐	☐	☐	☐	
POEDIL PEILOD	7	8	☐	☐	☐	12		☐	30	☐	☐	☐	☐	
TILCAMLE BNDO	4	16	10	31	25	☐	☐	☐		☐	☐	19	☐	
RIELURLEMTOCNA	2	☐	21	☐	☐	26	☐	☐	☐	☐	22	☐	27	☐
NAV RED SALWA	☐	24	☐		15	☐	23		☐	3	20	☐	17	
REENOCSAN	14	☐	9	☐	6	☐	☐	☐	28					
LALOOPETR	13	32	☐	☐	18	☐	5	☐	☐					

1 2 3 4 5 6 7

8 9   10 11 12

11 13 14 15 16 17 18

19 20 21 22 23 24 25

26 27 10 28 29 30 31 32

CRYPTOGRAM

Below is a phrase about bonding. Try to find out the whole phrase with the given clues.

A	B	C	D	E	F	G	H	I	J	K	L	M	N	O	P	Q	R	S	T	U	V	W	X	Y	Z
25					9			14							5		23	16						13	

14 19 6 10 23 3 18 2 10 4 26 2 25 23   9 18 23 4 10 16   25 9 9 10 4 6

6 22 10   5 22 13 16 14 4 25 2   5 23 18 5 10 23 6 14 10 16   18 9

6 22 10   16 26 12 16 6 25 19 4 10 16

GLOSSARY
ANSWERS
INDEX

CHEMICAL BONDS

GLOSSARY

Allotropes : The existence of more than one physical form of an element.

Atomic radius : The distance from the nucleus to the outer most electron orbital in an atom.

Bond length : The distance between the nuclei of two bonding atoms in a molecule.

Chemical bond : The attractive force that holds atoms together in a compound.

Covalent bond : A bond formed by the sharing of electrons between two non-metal atoms.

Dipole–Dipole forces : The attractions of the positive and negative poles of molecules are called dipole–dipole forces.

Dot representation : In this representation, the valence electrons are shown around the symbols as dots.

Double bond : A covalent bond in which two pairs of electrons are shared.

Electronegativity : A measure of the tendency of an atom to attract bonding electrons.

Electron configuration : The arrangement of electrons in atomic orbitals.

Ground state : The lowest energy state of an atom, molecule or ion.

Hybridization : The combination of orbitals from different energy levels to form new orbitals all with the same energy.

Hydrogen bond : An extra strong dipole-dipole attraction that occurs between molecules in which hydrogen is covalently bonded to the electronegative elements N, O and F.

Intermolecular forces : Attractive forces between neighboring molecules.

Ionic bond : A bond formed by the complete transfer of electrons between metal and nonmetal atoms.

Ionization energy : The energy required to remove an electron from a free atom or ion in the gaseous state.

Lewis structure : The structural formula drawn with Lewis symbols that shows the valence electrons using dots.

Line representation : Each line represents one bond or a lone pair of electrons.

Lone pair : A pair of electrons found in the valance shell of an atom that is not shared with another atom.

Metallic bond : The attractive force that holds metal atoms together.

Network solid : Covalent substances whose atoms are bonded together with a network structure.

Non–polar covalent bond : A covalent bond that is formed between two atoms with the same electronegativity values.

Octet rule : The rule that states an atom tends to lose or gain electrons until it has eight electrons in its valence shell.

Orbital representation : Orbital representation showing atomic orbitals in which the electrons are indicated as paired or unpaired.

Pi (π) bond : A bond that results from the sideways overlap of a pair of p orbitals.

Polar covalent bond : A covalent bond, that is formed between two atoms with different electronegativities.

Resonance structure : If the valence electrons in a molecule are capable of several alternative arrangements which differ only a small amount in energy each arrangement is called a resonance structure.

Sigma (σ) bond : A bond formed by the end to end overlap of pure or hybridized atomic orbitals.

sp hybrid orbitals : Hybrid orbitals formed by the combination of one s and one p atomic orbital.

sp^2 hybrid orbitals : Hybrid orbitals formed by combination of one s and two p atomic orbitals.

sp^3 hybrid orbitals : Hybrid orbitals formed by combination of one s and three p atomic orbitals.

Triple bond : A bond in which three pairs of electrons are shared.

Valence electron : The electrons of an atom which are found in the outer most energy level.

Van der Waals forces : The intermolecular forces between nonpolar molecules in the liquid and solid state.

ANSWERS

SUPPLEMENTARY QUESTIONS

Chapter_1 CHEMICAL BONDS

1. F > N > P > Si

4. **a.** N^{3-} : $1s^2\ 2s^2\ 2p^6$
c. S^{2-} : $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$
e. Ag^+ : $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^0\ 4d^{10}$

5.

a. 

c. 

e. 

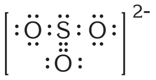
6.

a. 

c. 

e. 

g. 

i. 

8. Na_2O, KCl, $CaBr_2$

9. KF > AlF_3 > MgO > NaCl > CaS

13. **a.** Ionic bond
c. Polar covalent bond
e. Both ionic and covalent

22. **a.** 1 σ, 1 π
c. 1 σ, 2 π
e. 3 σ, 2 π

Chapter_2 BONDS IN LIQUIDS AND SOLIDS

2. Li > Na > K > Rb > Cs > Fr

3. II > I > III

4. KF ; KCl

11. b > c > a

13. **a.** C_5H_{12} **c.** S_8 **e.** diamond **g.** SiO_2

14. **a.** I_2 **c.** N_2 **e.** SCl_6 **g.** HCl **k.** Na_2S

15. **a.** metallic bond **c.** ionic bond **e.** dipole-dipole
g. hydrogen bond **i.** ionic bond **k.** metallic bond

16. Ne, CH_4, CO_2

MULTIPLE CHOICE

Chapter_1 CHEMICAL BONDS

1.	A	**5.**	E	**9.**	D	**13.**	B	**17.**	B
2.	E	**6.**	C	**10.**	B	**14.**	D	**18.**	D
3.	C	**7.**	E	**11.**	C	**15.**	B		
4.	D	**8.**	C	**12.**	A	**16.**	C		

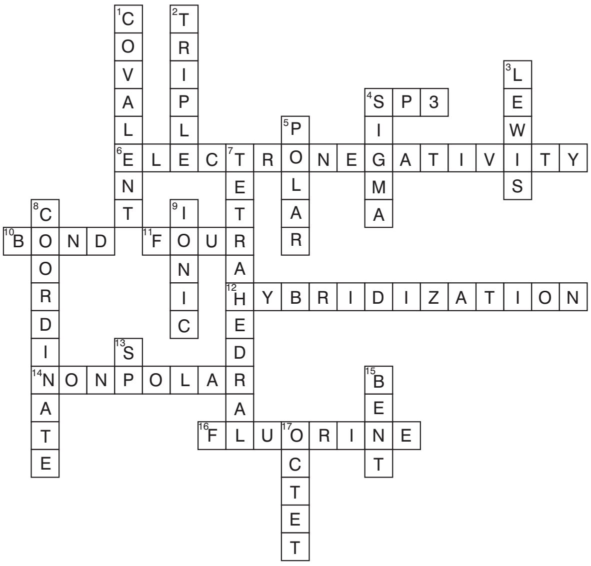
Chapter_2 BONDS IN LIQUIDS AND SOLIDS

1.	E	**5.**	E	**9.**	E	**13.**	B	**17.**	B
2.	A	**6.**	C	**10.**	C	**14.**	D	**18.**	B
3.	D	**7.**	C	**11.**	D	**15.**	E		
4.	D	**8.**	C	**12.**	D	**16.**	B		

PUZZLE

Chapter_1 CHEMICAL BONDS

CRISS – CROSS PUZZLE

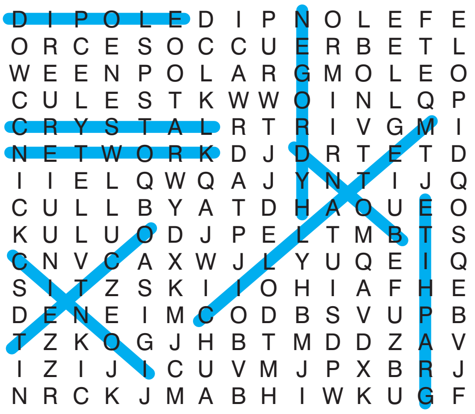


CRYPTOGRAM

CHEMICAL BOND IS THE ATTRACTIONAL FORCE THAT HOLDS ATOMS TOGETHER IN COMPOUNDS.

Chapter_2 BONDS IN LIQUIDS AND SOLIDS

WORD SEARCH



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W	E	E	N	P	O	L	A	R	G	M	O	L	E	O
C	U	L	E	S	T	K	W	W	O	I	N	L	Q	P
C	R	Y	S	T	A	L	R	T	R	I	V	G	M	I
N	E	T	W	O	R	K	D	J	D	R	T	E	T	D
I	I	E	L	Q	W	Q	A	J	Y	N	T	I	J	Q
C	U	L	L	B	Y	A	T	D	H	A	O	U	E	O
K	U	L	U	O	D	J	P	E	L	T	M	B	T	S
C	N	V	C	A	X	W	J	L	Y	U	Q	E	I	Q
S	I	T	Z	S	K	I	I	O	H	I	A	F	H	E
D	E	N	E	I	M	C	O	D	B	S	V	U	P	B
T	Z	K	O	G	J	H	B	T	M	D	D	Z	A	V
I	Z	I	J	I	C	U	V	M	J	P	X	B	R	J
N	R	C	K	J	M	A	B	H	I	W	K	U	G	F

DOUBLE PUZZLE

HYDROGEN BOND

DIPOLE DIPOLE

METALLIC BOND

INTERMOLECULAR

VAN DER WAALS

RESONANCE

ALLOTROPE

Secret Message

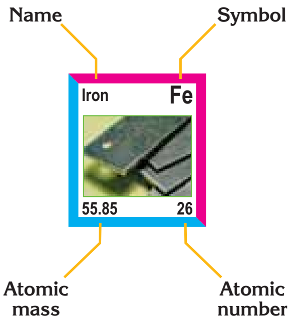
DIAMOND
IS THE
HARDEST
NATURAL
MATERIAL

CRYPTOGRAM

INTERMOLECULAR FORCES AFFECT THE PHYSICAL PROPERTIES OF THE SUBSTANCES.

INDEX

PERIODIC



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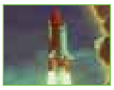
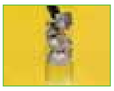
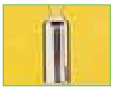
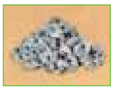
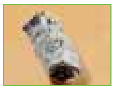
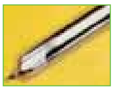
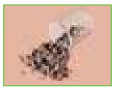
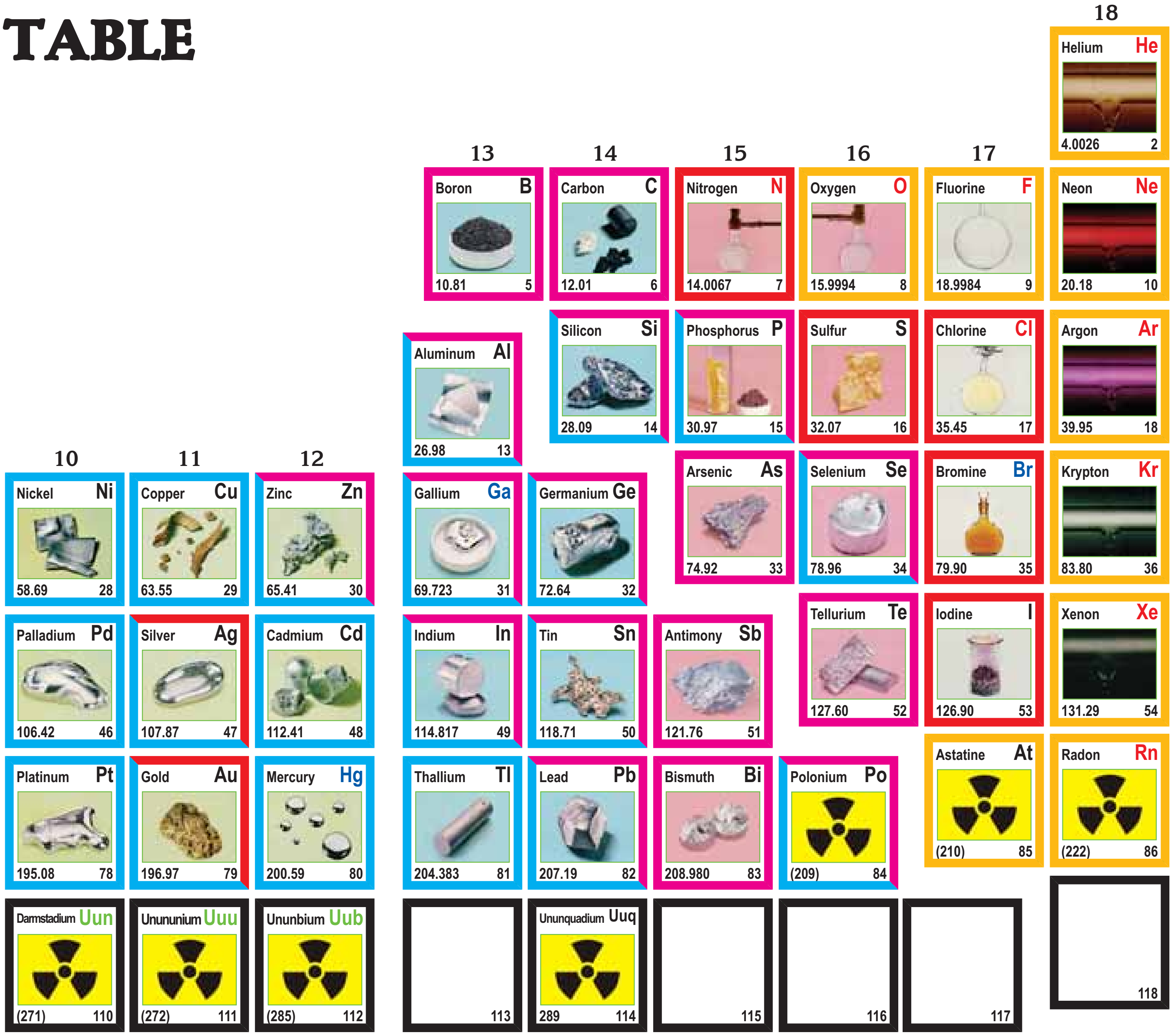
Hydrogen H  1.01 1									
Lithium Li  6.94 3	Beryllium Be  9.01 4								
Sodium Na  22.99 11	Magnesium Mg  24.31 12								
Potassium K  39.10 19	Calcium Ca  40.08 20		Scandium Sc  44.96 21	Titanium Ti  47.88 22	Vanadium V  50.94 23	Chromium Cr  52.00 24	Manganese Mn  54.94 25	Iron Fe  55.85 26	Cobalt Co  58.93 27
Rubidium Rb  85.47 37	Strontium Sr  87.62 38		Yttrium Y  88.91 39	Zirconium Zr  91.22 40	Niobium Nb  92.91 41	Molybdenum Mo  95.94 42	Technetium Tc ☢ (98) 43	Ruthenium Ru  101.07 44	Rhodium Rh  102.91 45
Cesium Cs  132.91 55	Barium Ba  137.33 56	Lanthanides * 57-70	Lutetium Lu  174.97 71	Hafnium Hf  178.49 72	Tantalum Ta  180.95 73	Tungsten W  183.84 74	Rhenium Re  186.21 75	Osmium Os  190.2 76	Iridium Ir  192.22 77
Francium Fr ☢ {223} 87	Radium Ra ☢ (226) 88	Actinides ** 89-102	Lawrencium Lr ☢ (262) 103	Rutherfordium Rf ☢ (261) 104	Dubnium Db ☢ (262) 105	Seaborgium Sg ☢ (266) 106	Bohrium Bh ☢ (264) 107	Hassium Hs ☢ (269) 108	Meitnerium Mt ☢ (268) 109

* Lanthanides	Lanthanum La  138.91 57	Cerium Ce  140.12 58	Praseodymium Pr  140.91 59	Neodymium Nd  144.24 60	Promethium Pm ☢ (145) 61	Samarium Sm  150.36 62
** Actinides	Actinium Ac ☢ (227.03) 89	Thorium Th ☢ 232.04 90	Protactinium Pa ☢ 231.04 91	Uranium U ☢ 238.03 92	Neptunium Np ☢ (237.05) 93	Plutonium Pu ☢ (244) 94

TABLE



10	11	12	13	14	15	16	17	18
								Helium He 4.0026 2
			Boron B 10.81 5	Carbon C 12.01 6	Nitrogen N 14.0067 7	Oxygen O 15.9994 8	Fluorine F 18.9984 9	Neon Ne 20.18 10
			Aluminum Al 26.98 13	Silicon Si 28.09 14	Phosphorus P 30.97 15	Sulfur S 32.07 16	Chlorine Cl 35.45 17	Argon Ar 39.95 18
Nickel Ni 58.69 28	Copper Cu 63.55 29	Zinc Zn 65.41 30	Gallium Ga 69.723 31	Germanium Ge 72.64 32	Arsenic As 74.92 33	Selenium Se 78.96 34	Bromine Br 79.90 35	Krypton Kr 83.80 36
Palladium Pd 106.42 46	Silver Ag 107.87 47	Cadmium Cd 112.41 48	Indium In 114.817 49	Tin Sn 118.71 50	Antimony Sb 121.76 51	Tellurium Te 127.60 52	Iodine I 126.90 53	Xenon Xe 131.29 54
Platinum Pt 195.08 78	Gold Au 196.97 79	Mercury Hg 200.59 80	Thallium Tl 204.383 81	Lead Pb 207.19 82	Bismuth Bi 208.980 83	Polonium Po (209) 84	Astatine At (210) 85	Radon Rn (222) 86
Darmstadium Uun (271) 110	Unununium Uuu (272) 111	Ununbium Uub (285) 112	113	Ununquadium Uuq 289 114	115	116	117	118

Europium Eu 151.96 63	Gadolinium Gd 157.25 64	Terbium Tb 158.92 65	Dysprosium Dy 162.50 66	Holmium Ho 164.93 67	Erbium Er 167.26 68	Thulium Tm 168.93 69	Ytterbium Yb 173.04 70
Americium Am (243) 95	Curium Cm (247) 96	Berkelium Bk (247) 97	Californium Cf (251) 98	Einsteinium Es (252) 99	Fermium Fm (257) 100	Mendelevium Md (258) 101	Nobelium No (259) 102