Stop Faking It!

Finally Understanding Science So You Can Teach It

MORE CHEMISTRY BASICS

BY WILLIAM C. ROBERTSON, PHD
IN CONSULTATION WITH MICHAEL S. KRALIK, PHD,
AND ANN CUTLER, PHD
ILLUSTRATED BY BRIAN DISKIN

NSTApress
National Science Teachers Association

Stop Faking It!

Finally Understanding Science So You Can Teach It

MORE CHEMISTRY BASICS

Stop Faking It!

Finally Understanding Science So You Can Teach It

MORE CHEMISTRY BASICS

By William C. Robertson, PhD
In consultation with Michael S. Kralik, PhD,
and Ann Cutler, PhD
Illustrated by Brian Diskin

NSTApress
National Science Teachers Association

Arlington, Virginia

Copyright © 2010 NSTA. All rights reserved. For more information, go to www.nsta.org/permissions.

NSTApress®
National Science Teachers Association

Claire Reinburg, Director
Jennifer Horak, Managing Editor
Andrew Cocke, Senior Editor
Judy Cusick, Senior Editor
Wendy Rubin, Associate Editor
Amy America, Book Acquisition Coordinator

ART AND DESIGN
Will Thomas Jr., Director

PRINTING AND PRODUCTION
Catherine Lorrain, Director
Jack Parker, Electronic Prepress Technician

SCILINKS
Tyson Brown, Director
Virginie Chokouanga, Customer Service and Database Coordinator

NATIONAL SCIENCE TEACHERS ASSOCIATION
Francis Q. Eberle, PhD, Executive Director
David Beacom, Publisher

Copyright © 2010 by the National Science Teachers Association.
All rights reserved. Printed in the United States of America.
13 12 11 10 4 3 2 1

Library of Congress Cataloging-in-Publication Data
Robertson, William C.
More chemistry basics / William C. Robertson ; in consultation with Michael S. Kralik and Ann Cutler; illustrated by Brian Diskin.
p. cm. -- (Stop faking it!)
Includes index.
ISBN 978-1-933531-47-2
1. Chemistry--Popular works. I. Title. II. Title: Chemistry basics.
QD37.R534 2010
540--dc22
2009052866
e-ISBN 978-1-936137-74-9

NSTA is committed to publishing material that promotes the best in inquiry-based science education. However, conditions of actual use may vary, and the safety procedures and practices described in this book are intended to serve only as a guide. Additional precautionary measures may be required. NSTA and the authors do not warrant or represent that the procedures and practices in this book meet any safety code or standard of federal, state, or local regulations. NSTA and the authors disclaim any liability for personal injury or damage to property arising out of or relating to the use of this book, including any of the recommendations, instructions, or materials contained therein.

PERMISSIONS
You may photocopy, print, or e-mail up to five copies of an NSTA book chapter for personal use only; this does not include display or promotional use. Elementary, middle, and high school teachers *only* may reproduce a single NSTA book chapter for classroom or noncommercial, professional-development use only. For permission to photocopy or use material electronically from this NSTA Press book, please contact the Copyright Clearance Center (CCC) (*www.copyright.com*; 978-750-8400). Please access *www.nsta.org/permissions* for further information about NSTA's rights and permissions policies.

SCILINKS THE WORLD'S A CLICK AWAY
Featuring sciLINKS®—a new way of connecting text and the Internet. Up-to-the-minute online content, classroom ideas, and other materials are just a click away.

Contents

Preface vii

Safety Note ix

How to Get Special Materials x

About the Author, Illustrator, and Consultants xi

Chapter 1 Déjà Review 1

Chapter 2 Dynamic Atoms 23

Chapter 3 The Name's Bond ... Pi Bond 43

Chapter 4 Special Reactions 67

Chapter 5 Electro-Luminescence 89

Chapter 6 Half a Life Is Better Than None 107

Chapter 7 A Little Organic 123

Glossary 135

Index 145

Preface

Prior to the publication of this book, teachers often asked me whether the book would be more advanced than the first book, *Chemistry Basics*. The answer to that question is yes and no. First, this book builds on concepts discussed in the first book. Although the first chapter of this book is a review of the concepts covered in the first book, the review will not suffice if you are not somewhat familiar with those concepts. So, you need to know a few things about basic chemistry before tackling this book. Second, there are a number of topics in this book that might be considered advanced chemistry, such as the uncertainty principle in quantum mechanics, the formation of different kinds of bonds, and oxidation-reduction reactions. Many of the reviewers of the book said things such as, "The teachers in your audience will never teach some of these concepts to their students. The students won't see some of these concepts until high school or even college." That might be true, but I believe that your understanding of the so-called higher-level concepts will enhance your teaching. When little Susie says, "These orbits of electrons aren't correct; my mom says we can't really know what electrons are doing," what will be your response? It could be, "We don't discuss that in sixth grade—you'll learn that in high school." But if you actually know what Susie is talking about, you have a choice. You can discuss it with her individually, or you might even decide that the subject will be great for a class discussion. If you don't know what Susie is talking about, you have no choice and no control over the situation. You need to ask yourself whether you want the content of your curriculum to be in charge of what happens, or if you want your knowledge of the subject matter to be in charge of what happens. I prefer the latter. Also, it is empowering to know more than you will ever need to teach the students. It gives you a comfort level and just might affect how you address certain topics. Teaching is more fun when you are secure in the depth of your understanding.

One more thing on this topic. There is a lot of tradition in the teaching of science. You simply don't address certain concepts until you get to a particular grade level. This implies that some ideas are much more difficult to understand than others. I strongly believe, however, that anyone can understand, say, the basics of quantum mechanics. Yes, a complete mathematical treatment of the subject requires a sophistication with math, but the basic concepts are not that difficult to understand. You can avoid the math complexities (as I do, for the most part, in this book) and still grasp the major concepts. As someone (I forget who) said, you can teach anyone of any age just about any concept, as long as you do so in an intellectually honest manner. This means that anyone can learn any concept, but you must take care not to alter the concept in the process to the point that you are teaching lies, which unfortunately happens too often when we try to "dumb down" certain material. Suffice to say that I try to remain intellectually honest in everything I write.

Content aside, this book is arranged the same as all the other books in the *Stop Faking It!* series. I ask you to do a few simple activities prior to the explanation of any concepts. We understand concepts better when we can anchor them to concrete experiences, and that's the purpose of these sections. It might be tempting to just read through the Things To Do sections, but I strongly encourage you to do the activities. They could make the difference between memorizing concepts and truly understanding them. Memorization fades, but understanding lasts.

Finally, I want to be clear that this is not a comprehensive chemistry book, even when combined with the first chemistry book in the series. For example, I do not deal with equilibrium in chemical equations or solubility of solids immersed in solutions, both traditional chemistry concepts. I also do not do justice to the vast field of organic chemistry. This book is not a textbook, and none of the books in the *SFI* series are intended to be comprehensive textbooks. My goal is to help you obtain a deep understanding of a number of basic chemistry concepts. I hope that with this understanding, you will know what those short descriptions of content in activity books are all about and will have the basis for knowing what's going on in a textbook. Also, although there are activities in this book, the activities are there to help you understand concepts and are not necessarily intended as classroom activities. That's not to say that you can't use some of these activities in the classroom, but rather that you should consider safety and appropriateness before using the activities with your students.

Safety Note

Though the activities in this volume don't require anything more volatile than household vinegar, safety should always be in the forefront of the mind of every teacher. (This is not intended as a book of classroom activities, by the way. Rather, the activities are designed to enhance your understanding of the subject before you get into the classroom.) Your individual school, or possibly the school system of which your school is a part, likely has rules and procedures for classroom and laboratory safety.

You can also find specific guidelines for the safe storage, use, and disposal of thousands of types of chemical products in the Material Safety Data Sheets (MSDS). Start with *http://www.ilpi.com/msds/#Internet*. This site links to dozens of free searchable databases, including those of top American and European universities.

NSTA has also published several award-winning titles covering the safety theme at all school levels. For the elementary level, there's *Exploring Safely: A Guide for Elementary Teachers* and the *Safety in the Elementary Science Classroom* flipchart. Middle school-level offerings are *Inquiring Safely: A Guide for Middle School Teachers* and the *Safety in the Middle School Science Classroom* flipchart. Finally, there's *Investigating Safely: A Guide for High School Teachers*.

How to Get Special Materials

You can visit the following websites to get supplies for chemistry activities.

Arbor Scientific. *www.arborsci.com/SearchResult.aspx?CategoryID=4*

Carolina. *www.carolina.com*

Edmund Scientific. *http://scientificsonline.com*

Educational Innovations. *www.teachersource.com/Chemistry.aspx*

Hi-Valley Chemical. *www.hvchemical.com*
(*Note:* Although other suppliers require that you are affiliated with a school or other institution, Hi-Valley will sell supplies to private individuals—which is important for home schoolers or others who are not affiliated with an institution.)

Sargent-Welch. *http://sargentwelch.com/chemicals/c/4749*

Science Kit. *http://sciencekit.com*

Ward's Natural Science. *http://wardsci.com/category.asp_Q_c_E_1251*

About the Author

As the author of NSTA Press's *Stop Faking It!* series, Bill Robertson believes science can be both accessible and fun—if it's presented so that people can readily understand it. Robertson is a science education writer, reviews and edits science materials, and frequently conducts inservice teacher workshops as well as seminars at NSTA conferences. Bill has published research in cognitive science that addresses the difference between memorizing and understanding and how that affects learning. He has also taught college-level physics and math and developed K–12 science curricula, teacher materials, and award-winning science kits. He earned a master's degree in physics from the University of Illinois and a PhD in science education from the University of Colorado.

About the Illustrator

The recently-out-of-debt, soon-to-be-famous, humorous illustrator Brian Diskin grew up outside of Chicago. He graduated from Northern Illinois University with a degree in commercial illustration, after which he taught himself cartooning. His art has appeared in many books, including *The Beerbellie Diet* and *How a Real Locomotive Works*. You can also find his art in newspapers, on greeting cards, on T-shirts, and on refrigerators. At any given time he can be found teaching watercolors and cartooning, and hopefully working on his ever-expanding series of *Stop Faking It!* books. You can view his work at *www.briandiskin.com*.

About the Consultants

Michael Kralik received his PhD in chemistry from the University of Utah with postdoctorate studies in chemistry, pharmacology, and toxicology. He has been faculty at the university and has conducted many faculty and staff development seminars. Kralik has established product development and manufacturing operations domestically and internationally, and has directed the development of hundreds of products for major corporations in chemical, medical, pharmaceutical, and electronics industries. He has developed K–12 science curricula, teacher inservice workshops, and many award-winning educational toys, games, and science kits.

Ann Cutler is the field editor for the *Journal of College Science Teaching,* a position she has held since 2006. She is an associate professor of chemistry at the University of Indianapolis in Indiana and has a doctorate in inorganic chemistry from Purdue University.

Acknowledgments

The *Stop Faking It!* series of books is produced by the NSTA Press: Claire Reinburg, director; Jennifer Horak, managing editor; Wendy Rubin, associate editor; Will Thomas Jr., art director; Catherine Lorrain, printing and production director; and Jack Parker, electronic prepress technician. Will Thomas Jr. designed the cover from an illustration provided by artist Brian Diskin, who also created the inside illustrations.

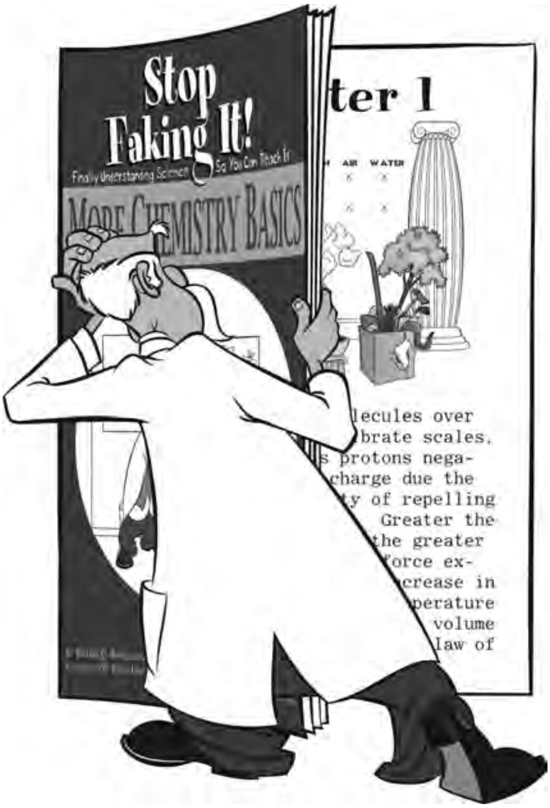
This book was reviewed by Dennis Huffman (Lanier Middle School, Texas), Sue Vogel (Thomas Jefferson High School, Iowa), and Peggy Carlisle (Pecan Park Elementary School, Mississippi).

Dedication

I dedicate this book to two people. The first is Michael Kralik, a good friend and valuable colleague for many years who has served as a consultant on both chemistry books. There's a saying in science education that all activities are recycled—there really isn't anything new. People who say that have never met Michael, who consistently comes up with inventive ways to introduce concepts. The second person is Ann Cutler, who was a reviewer for both chemistry books. I can honestly say that Ann has fundamentally changed the content of both books and, with this second book, saved me from embarrassing myself with hastily written, wrong explanations.

1 Chapter

Déjà Review



This is the second chemistry book in the *Stop Faking It!* series; the first one is called *Chemistry Basics*. This second book introduces new concepts and expands on many of the concepts presented in the first book, so it would help if you knew the contents of the first book before picking up this one. Now, I can't assume that you have already worked your way through the first book. It's also possible that you worked your way through that book but need a refresher on the contents. Hence, I'm beginning this book with a chapter that reviews the content of *Chemistry Basics*. In no way should you consider this chapter a substitute for going through the first book, though. If your knowledge of chemistry concepts is at all shaky, I strongly advise you to work through *Chemistry Basics* first.

Because this is a review chapter, it does not follow my usual format. Normally, I have you, the reader, perform a few simple activities before reading explanations of concepts. The explanations use the results of the activities to help you anchor the concepts to experiences, thus leading to a deeper understanding than would otherwise be possible. If I did the activity thing in this chapter, then this chapter would be as long as the first book, and that wouldn't make much sense. So, what follows here are explanations of concepts *sans* activities.

As such, this first chapter represents what you normally expect from a textbook or other science resource, which is a set of straightforward explanations. If you do own *Chemistry Basics,* comparing this chapter to that book will provide a clear example of what it's like to learn science concepts with and without use of the Learning Cycle.

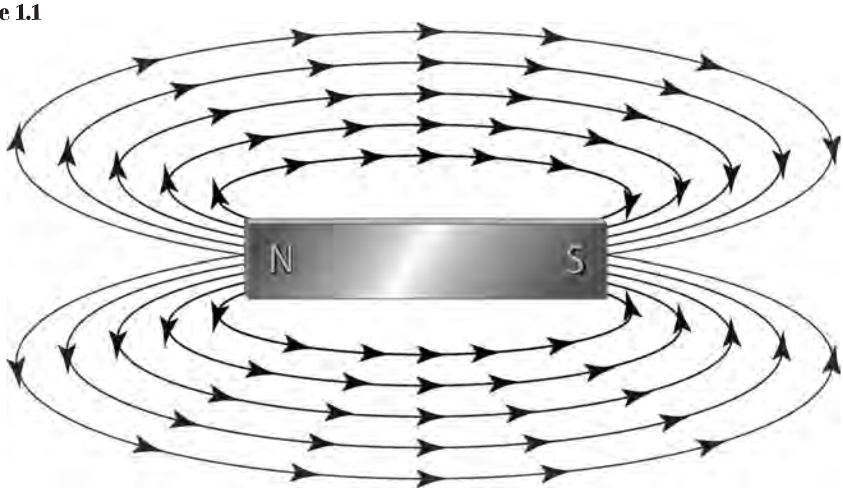
And yes, I realize that *déjà review* is the same as saying *re-review*. Makes sense to me, though, because there is a summary section at the end of each chapter in *Chemistry Basics*. You can consider a summary section a review, so this chapter counts as the second review of the material. And yes, I realize that most of you don't care whether or not I justify the chapter title!

Using models to explain observations

The primary building block of chemistry is **atomic theory**—our model of what tiny little things called atoms look like and how they behave, both alone and in interactions with other atoms. Using atoms to explain observations is typical of what scientists do—develop **models**. For example, two magnets exert forces on each other. These forces can be attractive or repulsive. To describe and explain such interactions, physicists invented things called **magnetic fields**. With rules for the magnetic field lines that magnets generate and for how magnets and other objects react when in the presence of magnetic field lines, you can understand all sorts of things magnetic in nature. Figure 1.1 shows the magnetic field lines associated with a bar magnet.[1]

SCILINKS
THE WORLD'S A CLICK AWAY
Topic: Atomic Theory
Go to: *www.scilinks.org*
Code: MCB001

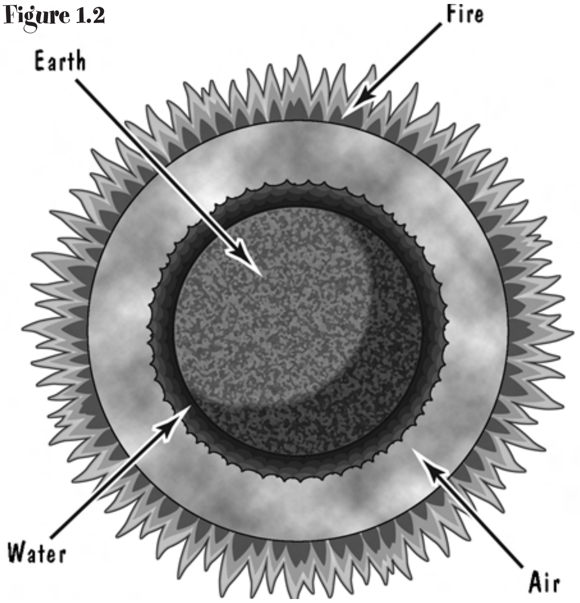
Figure 1.1



---

[1] The magnetic field lines shown in Figure 1.1 are just an example of the models that scientists use. I'm not expecting you to know what these field lines represent with respect to magnetic interactions.

Figure 1.2



The process of creating scientific models began long ago. One of the earliest models to explain the workings of the universe came from the early Greeks, who proposed that there are four main elements in the universe: earth, air, water, and fire. All things in the universe are composed of these elements in different proportions, and these four elements strive to attain their "proper" places in the universe, illustrated in Figure 1.2. You can explain the properties of objects and what happens when different substances interact simply by determining how much of each of the four basic elements the substances contain.

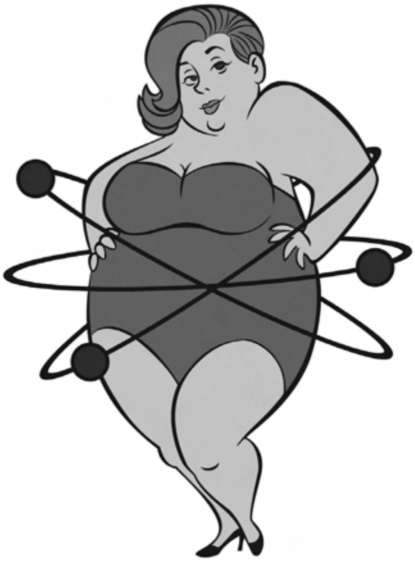
SCILINKS
THE WORLD'S A CLICK AWAY
Topic: Atomic Models
Go to: *www.scilinks.org*
Code: MCB002

It turns out you can use the model of earth, water, air, and fire to explain many occurrences. This model was so effective, in fact, that it's a wonder that the theory that atoms (tiny, indivisible, marblelike objects that make up everything in the universe) exist ever got a foothold in the scientific and philosophic communities. The theory of atoms was established, though, and over the span of hundreds of years the model of the atom underwent lots of changes and currently is a pretty sophisticated model. The next section explains our current model of the atom. I use the word *current* because it's a pretty good bet that our model of the atom a hundred years from now will look quite a bit different from our model of the atom today. Personally, I'm expecting a similar result when I look into a mirror a hundred years from now, assuming that living to be 160 years old will soon be commonplace.

Atomic fashions, or what not to wear if you want to keep up with the halogens

The rather obscure title of this section means that I intend to explain scientists' currently accepted model of what your typical atom looks like. Atoms are all different, but they have many characteristics in common. Most books address the model of the atom as it developed historically, and my first chemistry book in this series takes that approach. Here, though, it's probably best to just cut to the chase and present a modern picture of the atom.

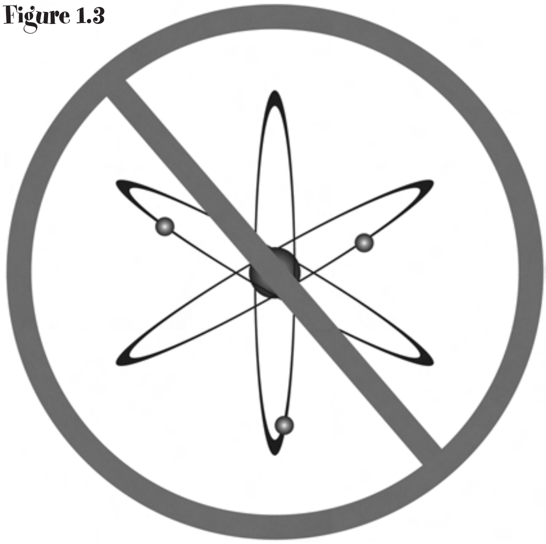
Atoms are composed of a positively charged, concentrated center called the **nucleus**, surrounded by negatively charged **electrons**. The nucleus contains positively charged **protons** and zero-charged **neutrons**. In electrically neutral atoms, there are equal numbers of protons and electrons. The number of neutrons in a nucleus varies. Hydrogen has no neutrons, and many of the lighter atoms have equal numbers of protons and neutrons. Even within one type of atom, the number of neutrons can vary. Atoms that have the same number of protons but different numbers of neutrons are called **isotopes**.


Atomic fashions

Before moving on, I want to show you a drawing of an atom that one often finds in elementary school textbooks and reference books for laypeople. Figure 1.3 shows protons and neutrons in the nucleus, surrounded by electrons moving in orbits not unlike the orbits of planets around the Sun.

Note the circle and slash through this drawing, indicating that it's wrong. Electrons do *not* orbit the nucleus like this. If they did, they would radiate energy away rapidly, resulting in a quick collapse of the universe. No collapsing universe that I notice, so why do books still use this outdated drawing? I've had people tell me it's useful for early grades because it's easier to understand than a more accurate drawing (which we'll get to later), but I'm not buying that. My position is that you can explain just about any concept to any age group in an intellectually honest way, which means you might need to

Figure 1.3



simplify things a bit, but you can stay true to the correct model. Orbits aren't true to the correct model, so I think an adjustment is in order. Unfortunately, the textbook companies don't ask for my advice much!

Anyway, let's move on. Protons and neutrons have approximately the same mass (we'll usually assume that the masses are equal), and both are much more massive than electrons. The mass of an electron is so small, in fact (over 1,000 times smaller than the mass of a proton), that it's common to ignore the electron mass when determining the mass of an atom. Electrons also are physically smaller than protons and neutrons—so small, in fact, that they don't have a measurable size.[2] Even though electrons are that tiny, they do help determine the overall size of an atom because they move around in a "cloud" that extends beyond the confines of the nucleus. The farther out the electron cloud extends, the larger the atom.

More details

We do know more about an atom than what makes it up. We have some idea of what the protons, neutrons, and electrons are doing. The protons and neutrons in the nucleus (Remember, the nucleus by far makes up most of the mass in an atom.) pretty much stay put in relation to the atom as a whole, even though all atoms in the universe are themselves moving at all times.[3] The electrons are a different story. As I mentioned, they don't move in circular orbits. So, what are the electrons doing? We know that they have only certain energies, with those energies being dictated by the type of atom you have. We speak of atoms as having **energy levels** or **energy shells** in which the electrons can reside. It's sort of like the electrons can sit in a football stadium but only in rows 2, 7, 15, and so on. The in-between rows (in-between energies in an atom) aren't available to the electrons. It's not that those energies don't exist but rather that they're off-limits for electrons in different atoms. This fits with the notion that electrons behave as bundles of waves (a premise of the mathematics of quantum mechanics) that are confined, in this case confined because of the electric attraction between the electrons

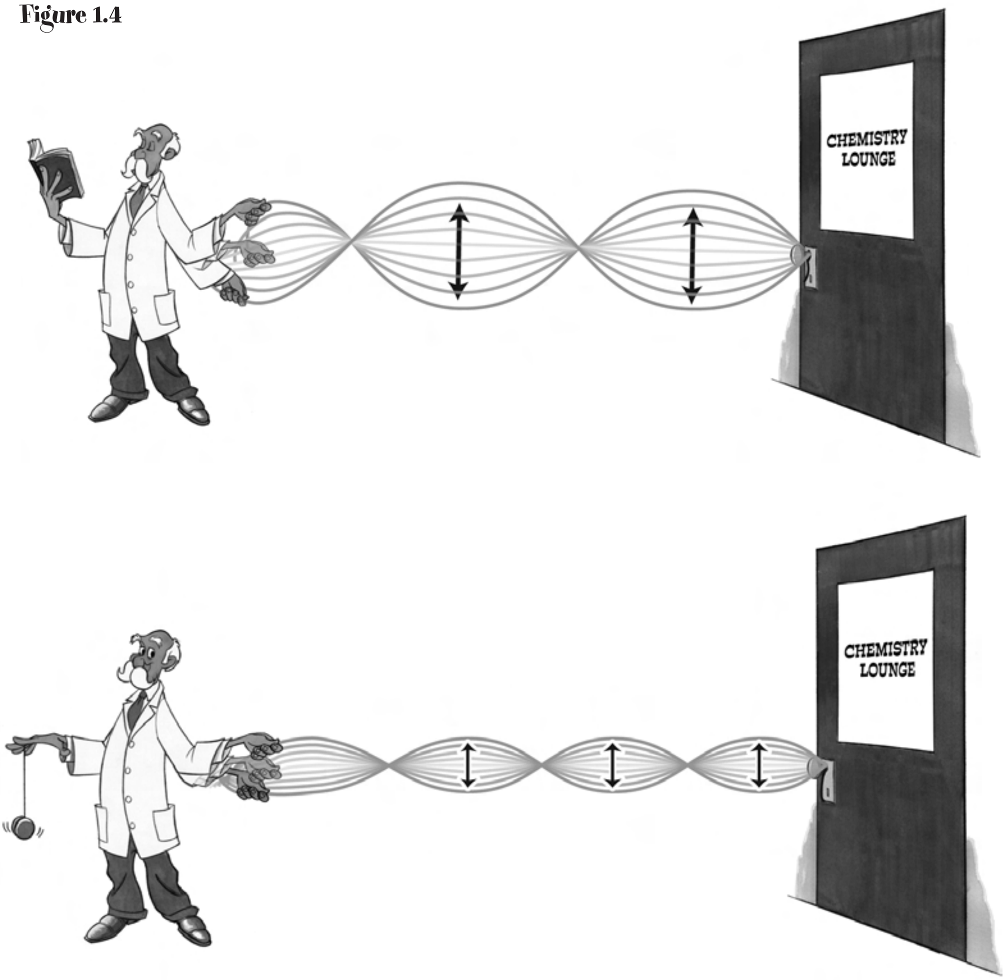
SCILINKS
THE WORLD'S A CLICK AWAY
Topic: Energy Levels
Go to: *www.scilinks.org*
Code: MCB003

---

[2] It might seem strange to you that something doesn't have a measurable size, but that's the case with electrons as well as with other subatomic objects in the universe. Scientists consider such objects to be point objects. And if the existence of point objects troubles you greatly, keep in mind one of my favorite phrases: All of science is made up. Mathematically, electrons act like they have zero size, so that's what we have to deal with, even if it doesn't seem to make sense.

[3] Here I refer to a random jiggling motion that characterizes all atoms, even those that make up solids. Even at the lowest temperatures possible, atoms still jiggle just a bit.

Figure 1.4



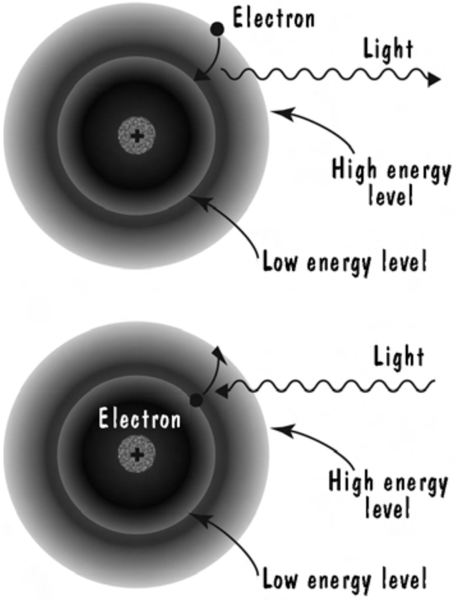
and the nucleus. When you confine waves, only certain "resonant" energies are allowed, as demonstrated using waves on a string shown in Figure 1.4.[4]

The electrons can only have certain energies, but they can jump from one energy level to the next. In the process, the atom either gives off or absorbs electromagnetic waves, some of which we know as light. Figure 1.5 shows an electron jumping between two energy levels in an atom. This drawing is a cross-section of the two lowest energy levels in an atom. These energy levels (orbitals) are spherical in shape and are known as the 1*s* and 2*s* orbitals.

---

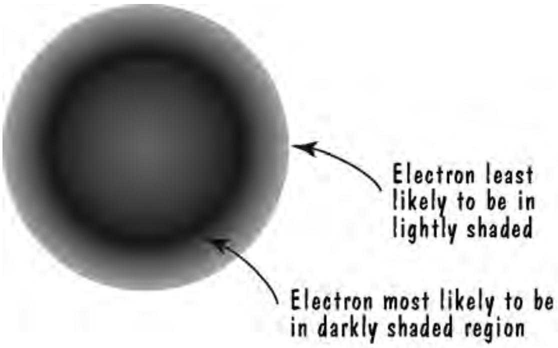
[4] Keep in mind that this is a review chapter. If this is your first introduction to the fact that electrons reside in distinct energy levels and why that is, then you really should pick up a more basic resource on this (say, my first chemistry book in the series ... hint, hint) to understand the topic better.

Figure 1.5



Okay, so we know that the electrons can have certain energies and jump from one energy level to another. What are they *doing* in the meantime? Well, we don't know for sure. The best we can come up with is a **probability distribution** for each electron, which is a mathematical description of the probability of finding the electron in any given place at any given time. Pictorially, we show this as shaded regions. Figure 1.6 shows a cross-section of the probability distribution for the lowest energy level in any atom (the entire lowest-energy probability distribution is spherical). The darker the shading, the more likely the electron is to be found at that location.[5]

Figure 1.6

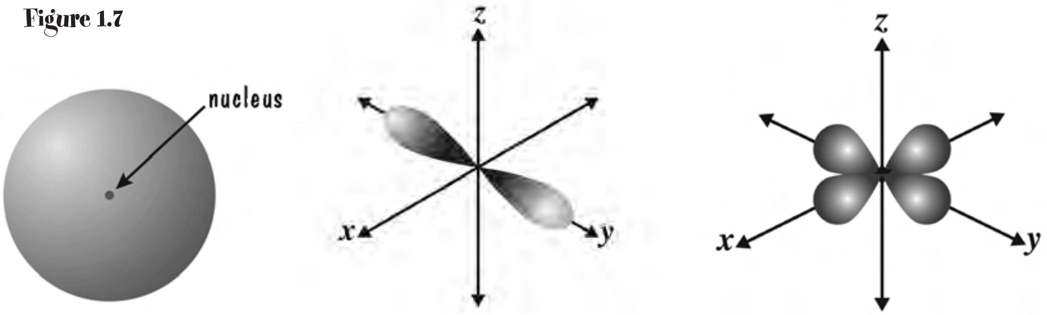


As you go from lower to higher energy levels in an atom, the shapes of the probability distributions change. Figure 1.7 (p. 8) shows the shapes of the distributions for three different energies of a major energy level within an atom. We call each of these distributions and parts of distributions **orbitals**, so it is common to refer to an electron as residing in "such and such an orbital" rather than give its precise energy. For example, we might talk about a 3*s* electron (3 referring to the main energy level and *s* to its sub-energy level, or orbital) or a 4*d* electron (again, the 4 is the main energy level and *d* is its sub-energy level or orbital).

---

[5] When I discuss orbitals later in the book, I'll explain why we can't know an electron's exact location.

Figure 1.7



The 1*s* energy level probability distribution

The 2*p* energy level probability distribution

The 3*d* energy level probability distribution

Low energy is best

One major principle that governs the universe is that systems (any specific object or collection of objects) tend toward configurations that have the lowest energy.[6] What that means is that objects in a system tend to spontaneously rearrange themselves to achieve a lower energy. For example, a rock on a cliff has more energy at the top of the cliff (farther from the Earth) than at the bottom (closer to the Earth), so its natural tendency is to fall to the bottom, where the energy is lower. Similarly, because negative electrons are attracted to the positive nucleus in an atom, the atom can have lower energy when the electrons are closer to the nucleus. Closer to the nucleus generally means lower energy levels for the electrons.[7] We could also say that electrons tend to drop to the lowest energy orbitals possible. Now, the lowest energy *possible* in a system of objects might not be the lowest energy *imaginable*. With our rock falling off a cliff, the rock might get stuck on a ledge halfway down the cliff; then the rock is at the lowest energy possible (the ledge stopped it) but not the lowest energy imaginable (the bottom of the cliff).

From rocks to electrons in atoms. The lowest energy level available to an electron might not be the lowest energy level imaginable of the atom. You see, each orbital in an atom can hold only so many electrons. The 1*s* orbital in an atom can only hold two electrons, so if you add an electron to an atom that already has two electrons, the new electron has to reside in the next highest energy level,

---

[6] This principle is not anything magical but rather just a convenient way to figure out what will happen in different situations. Systems tend toward lowest energy as a direct consequence of definitions of energy. Not gonna show you that right now.

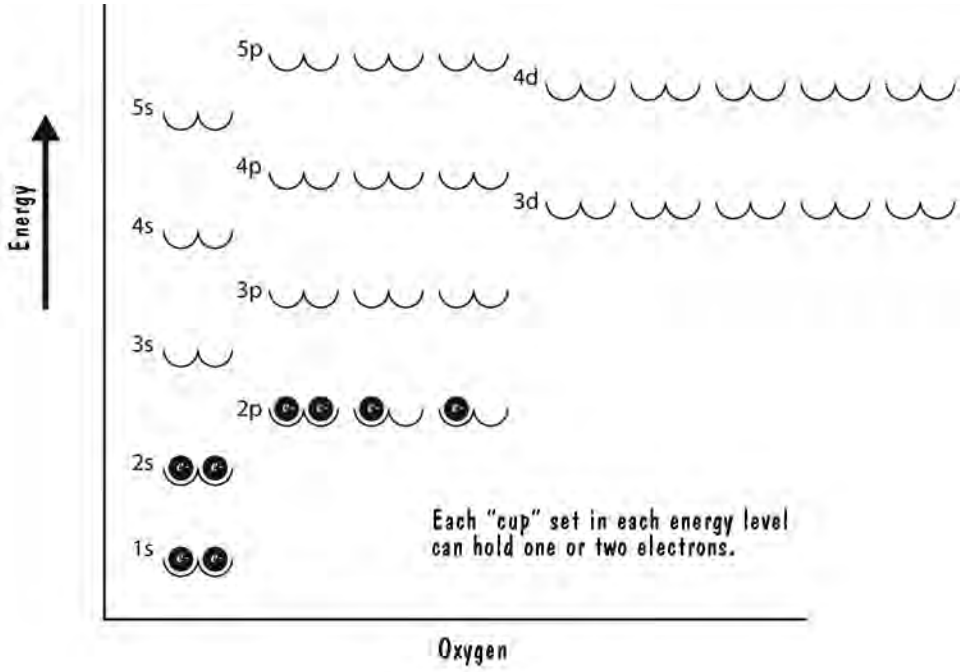
[7] Don't let Figure 1.7 confuse you about this idea. The drawings there reflect the shapes of the different orbitals but not necessarily their relative sizes. So, even though it looks like an electron in the *d* orbital (far-right drawing) might spend more time near the nucleus than an electron in the *s* orbital (far-left drawing), that's not the case.

which happens to be the 2*s* orbital. A good analogy is the game of Stadium Checkers (shown in Figure 1.8; search the internet for a description of the game) with the bottom holes plugged. You can drop marbles down toward the center by moving the rings. Once the lowest ring is full, new marbles have to reside in the next ring up. When that one's full, new marbles have to reside in the next ring up from that one.

Figure 1.8

So, here's our picture of what electrons are doing in an atom. We can draw orbitals, or probability distributions, that tell us where the electrons are *likely* to be. Electrons in different orbitals have different energies, and electrons tend to reside in the lowest energy orbital possible. We think of energy levels in an atom as existing even when they don't contain electrons, just as there are spaces available for marbles in Stadium Checkers even if the spaces don't contain a marble. This whole situation lets us draw energy level diagrams for atoms. The energy level diagram for oxygen is shown in Figure 1.9. Note that the electrons occupy the lowest energy levels possible, even though slots in higher energy

Figure 1.9



levels are available. Also note that in energy level diagrams the horizontal axis has no meaning other than to show different orbitals distinctly. The only direction to worry about is the vertical direction, which separates the electron positions by energy.

If you add energy to an atom (by heating it up, maybe), you can send electrons into higher energy slots. The electrons then spontaneously head back to the lower energy slots and emit electromagnetic waves, which might be in the visible part of the spectrum (light). Remember that we refer to different main energy levels as *shells*. It turns out that filled shells have lower energies than unfilled shells or orbitals. This is important when two or more atoms get together. For example, our oxygen atom would just love to grab a couple of electrons from another atom to fill up its 2*p* orbital, thus filling up its outer shell. The *way* in which atoms grab or give up electrons is the subject of bonds between atoms, covered later in this chapter.

The concept of filled and unfilled electron shells in atoms is extremely important for understanding how atoms behave around other atoms. Filled shells are usually quite stable and generally correspond to lower energies than unfilled shells. So in a sense, atoms "like" to have filled electron shells.[8] Different orbitals can accommodate different numbers of electrons, so what constitutes a filled shell depends on the orbitals involved. Filled *s* orbitals contain two electrons, filled *p* orbitals contain six electrons, and filled *d* orbitals contain ten electrons. See Table 1.1.

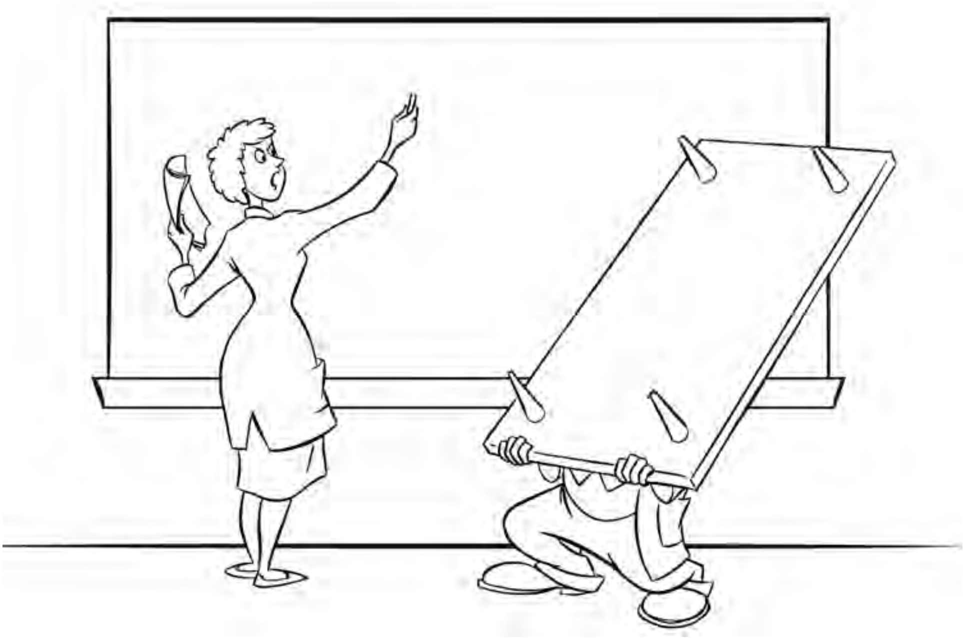
Table 1.1

Orbital	**Maximum number of electrons**
s	2
p	6
d	10
f	14

More on this as we discuss …

---

[8] Some people get all apoplectic over anthropomorphism (assigning human characteristics such as likes and dislikes to nonhuman animals or objects) in science. I don't. I say if it helps you understand, go for it. If anyone really thinks atoms think like humans, science might not be that person's thing.



"No, Professor. I said 'Periodic Table,' not 'periodical table.'"

The Periodic Table

And so we come to the one thing that causes an increase in blood pressure if one had a bad experience in chemistry—the Periodic Table (Figure 1.10, p. 12). Most likely that's because many of us were required to memorize some or most of this table in school. If you look at the table as an organizational chart or map, however, it's not very scary but actually quite useful. Each place in the table is reserved for a particular atom or element. An element is something composed of only one kind of atom, so for the purposes of studying the Periodic Table, we'll use the terms *element* and *atom* interchangeably. Thus, I might refer to an atom of carbon or the element carbon when talking about what occupies the sixth place in the Periodic Table.

Let's run through what all the numbers and such on this table mean. The letters stand for different elements (or atoms): H for hydrogen, He for helium, O for oxygen, Be for beryllium, Ag for silver, and on and on and on.[9]

---

[9] You might be troubled by the fact that element symbols sometimes don't seem to represent the element names, such as Ag for silver and W for tungsten. This is because contributions to the positions of elements on the Periodic Table have come from all over the world. The original German name for tungsten, for example, is *wolfram*. Hence, W represents tungsten in the table. The Latin name for silver is *argentum,* leading to the symbol Ag for silver. Further trivia: The Latin name for mercury is *hydroargentum* (leading to a symbol of Hg), which means liquid silver, which is what mercury looks like.

Figure 1.10

Periodic Table of Elements

1 H Hydrogen 1.00794																	2 He Helium 4.00260
3 Li Lithium 6.941	4 Be Berylium 9.01218											5 B Boron 10.81	6 C Carbon 12.011	7 N Nitrogen 14.0067	8 O Oxygen 15.9994	9 F Flourine 18.998403	10 Ne Neon 20.179
11 Na Sodium 22.98977	12 Mg Magnesium 24.305											13 Al Aluminum 26.98154	14 Si Silicon 28.0855	15 P Phosphorus 30.97376	16 S Sulfur 32.06	17 Cl Chlorine 35.453	18 Ar Argon 39.948
19 K Potassium 39.0983	20 Ca Calcium 40.08	21 Sc Scandium 44.9559	22 Ti Titanium 47.88	23 V Vanadium 50.9415	24 Cr Chromium 51.996	25 Mn Manganese 54.9380	26 Fe Iron 55.847	27 Co Cobalt 58.9332	28 Ni Nickel 58.69	29 Cu Copper 63.546	30 Zn Zinc 65.38	31 Ga Gallium 69.72	32 Ge Germanium 72.59	33 As Arsenic 74.9216	34 Se Selenium 78.96	35 Br Bromine 79.904	36 Kr Krypton 83.80
37 Rb Rubidium 85.467	38 Sr Strontium 87.62	39 Y Yttrium 88.9059	40 Zr Zirconium 91.22	41 Nb Niobium 92.9064	42 Mo Molybdenum 95.94	43 Tc Technetium (98)	44 Ru Ruthenium 101.07	45 Rh Rhodium 102.9055	46 Pd Palladium 106.42	47 Ag Silver 107.8682	48 Cd Cadmium 112.41	49 In Indium 114.82	50 Sn Tin 118.69	51 Sb Antimony 121.75	52 Te Tellurium 127.60	53 I Iodine 126.9045	54 Xe Xenon 131.29
55 Cs Cesium 132.9054	56 Ba Barium 137.33	57 La Lanthanum 138.9055	72 Hf Hafnium 178.49	73 Ta Tantalum 180.9479	74 W Tungsten 183.85	75 Re Rhenium 186.207	76 Os Osmium 190.2	77 Ir Iridium 192.2	78 Pt Platinum 195.08	79 Au Gold 196.9665	80 Hg Mercury 200.59	81 Tl Thallium 204.383	82 Pb Lead 207.2	83 Bi Bismuth 208.9804	84 Po Polonium (209)	85 At Astantine (210)	86 Rn Radon (222)
87 Fr Francium (223)	88 Ra Radium (226)	89 Ac Actinium (227)	104 Rf Rutherfordium (261)	105 Db Dubnium (262)	106 Sg Seaborgium (266)	107 Bh Bohrium (264)	108 Hs Hassium (269)	109 Mt Meitnerium (268)	110 Uun Ununnium	111 Uuu Unununium (285)	112 Uub Ununbium (284)	113 Uut Ununtrium (289)	114 Uuq Ununquadium (288)	115 Uup Ununpentium (292)	116 Uuh Ununhexium	117 Uus Ununseptium	118 Uuo Ununoctium

58 Ce Cerium 140.12	59 Pr Praseodymium 140.9077	60 Nd Neodymium 144.24	61 Pm Promethium (145)	62 Sm Samarium 150.36	63 Eu Europium 151.96	64 Gd Gadolinium 157.25	65 Tb Terbium 158.9254	66 Dy Dysprosium 162.50	67 Ho Holmium 164.9304	68 Er Erbium 167.26	69 Tm Thulium 168.9342	70 Yb Ytterbium 173.04	71 Lu Lutetium 174.967
90 Th Thorium 232.0381	91 Pa Protactinium 231.0359	92 U Uranium 238.0289	93 Np Neptunium 237.0482	94 Pu Plutonium (244)	95 Am Americium (243)	96 Cm Curium (247)	97 Bk Berkelium (247)	98 Cf Californium (251)	99 Es Einsteinium (252)	100 Fm Fermium (257)	101 Md Mendelevium (258)	102 No Nobelium (259)	103 Lr Lawrencium (260)

Topic: Periodic Table
Go to: *www.scilinks.org*
Code: MCB004

The main number on each element (1 for hydrogen, 2 for helium, 20 for calcium) is called the **atomic number**, which tells you how many protons are in the atom. The number of protons in an atom is the atom's main distinguishing characteristic. If you change the number of protons (which is not an easy thing to do), you have a different atom that corresponds to a different element. For a neutral atom, the atomic number also tells you the number of electrons in the atom. A neutral atom has no net charge because there is an equal number of positive charges (protons) and negative charges (electrons). **Ions**—atoms that have extra electrons or lack electrons—are still the same element because they have the same number of protons. The atomic number is always an integer—you can't have a fraction of a proton or an electron

The second number, at the bottom of each listed element in the Periodic Table, is called the **atomic mass**.[10] Atomic masses are not nice, round numbers but rather numbers such as 28.086 and 65.37. If you round off the atomic mass, the number you end up with tells you the total number of protons and neutrons in the most commonly occurring form of the atom. The atomic mass is measured in what are called **atomic mass units**. In this system, the mass of a proton is equal to exactly 1 and the mass of a neutron is equal to exactly 1.[11] The mass of an electron is really small compared to the masses of protons and neutrons, so we can ignore the electron's mass for most purposes in chemistry. So the electrons are still there in our atoms—we just don't count their mass.

An important thing to remember: Atoms with different numbers of protons are different elements. Different numbers of electrons or neutrons, however, do not necessarily indicate different elements.

Now take a look at hydrogen in the Periodic Table. Hydrogen contains 1 proton and 1 electron, with zero neutrons. Why, then, isn't the atomic mass of hydrogen equal to exactly 1? The reason is that in nature each element can have isotopes. For example, most of the time hydrogen has 1 proton and zero neutrons, but hydrogen can also have 1 neutron (this is called deuterium) or 2 neutrons (this is called tritium) and still be hydrogen. The atomic masses of elements are not integers because of the existence of isotopes. Atomic masses

---

[10] Older texts and even recent publications use the term *atomic weight* instead of *atomic mass*. The proper term is *atomic mass*, but for years the convention was to use *atomic weight*, and old habits die hard.

[11] We use atomic mass units instead of actual masses for simplicity. Also, the mass of a proton is different from the mass of a neutron, but the difference is small enough that we don't have to worry about it here.

are *averages* of the atomic masses that naturally occur in nature, and virtually all atoms can exist in isotope form.

As we move from left to right across the table, and as we form new rows, we encounter atoms that have one more proton than the previous atom. With neutral atoms, that also means that each successive atom has one more electron than the previous atom (that statement is not necessarily true for ions!). So, while protons determine an atom's identity, we sometimes focus just on what's happening to the electrons as we go from element to element.

Recall that as you add electrons to energy levels, the electrons fill the lowest energy levels first. Once lower energy levels are filled (remember that energy levels can hold only so many electrons), electrons fill the higher energy levels, just as in Stadium Checkers. So, as you move from left to right across the Periodic Table, you are adding protons and electrons[12] to atoms. As this happens, the electrons—in cooperation with the nuclei of atoms—obey the simple principle that systems naturally tend to be in their lowest energy.

Let's look at this pattern in more detail. In hydrogen, as in other atoms, the lowest energy level is 1*s*. Two electrons can reside in that energy level. When you add a proton and two neutrons to hydrogen, you get helium. You also get another electron, which fills the 1*s* energy level. I mentioned before that atoms with nothing but filled energy levels are extremely stable and represent low-energy situations. If you want to add an electron to a filled energy level, clearly that's a difficult thing to do. It is also difficult to remove an electron from an entirely filled energy level.

Noble gas

Well, we have helium with a filled 1*s* energy level. Helium is a **noble gas**. Noble gases are also called **inert gases**. The word *inert* means unresponsive or inactive, and that describes helium. Helium just plain doesn't interact much with other atoms.

SCILINKS® THE WORLD'S A CLICK AWAY

Topic: Inert Gases
Go to: *www.scilinks.org*
Code: MCB005

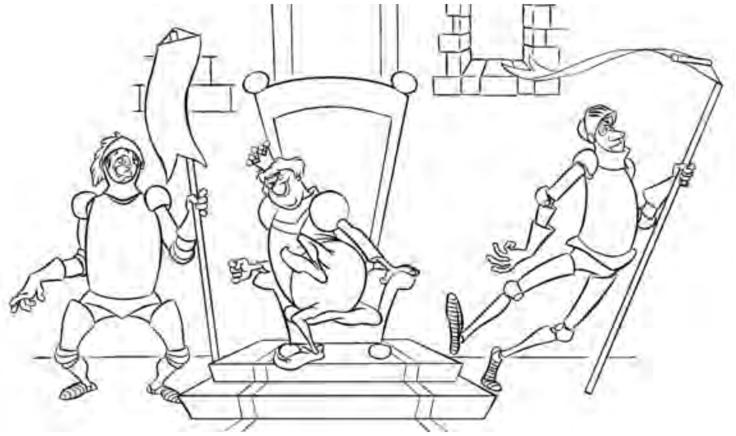
Before continuing, recall that scientists refer to the energy levels in atoms as **shells**. Outer shells are higher energy levels and inner shells are lower energy levels. This terminology no doubt dates back to when people thought of electrons as moving in circular orbits around the nucleus, which of course we now know they don't. Electrons in an outermost shell are known as **valence electrons**. We only consider *s* and *p* electrons when counting valence electrons.[13] We

---

[12] Keep in mind that we're not physically adding protons and electrons but describing a pattern in the table.

[13] Electrons in *d* and *f* orbitals generally aren't considered valence electrons because, as explained in a few pages, adding electrons to *d* and *f* orbitals occurs in inner shells.

would say, then, that hydrogen has one valence electron and helium has two valence electrons.[14]



Noble Gas

Back to the Periodic Table. In the second row, we start with lithium. It has a completely filled 1*s* energy level, plus an extra electron that hangs out in the 2*s* energy level. With an open slot in that energy level, lithium readily interacts with other atoms. The next atom in the second row is beryllium, which has a filled 2*s* energy level and, thus, two valence electrons. Shouldn't be a surprise that beryllium isn't quite as anxious as lithium to interact with other atoms because it has a filled 2*s* energy level, which is more stable than a half-filled 2*s* energy level.

Next we jump across a gap to the element boron. The reason we jump a gap will be clear later. After you jump the gap, you will notice that there are five elements before you get to the final atom on the right, which is neon. What's happening as you move through those five atoms is that electrons are filling the 2*p* energy level. By the time you get to neon, the 2*p* energy level is filled. Neon has a filled "shell" of eight electrons (recall that the *s* energy level holds up to two electrons and the *p* energy level can hold up to six electrons), and so, like helium, neon is an inert gas. Filled shells with eight valence electrons don't interact much with other atoms, so neon doesn't take part in many chemical reactions. You might expect that those five atoms before you get to argon—boron, carbon, nitrogen, oxygen, and fluorine—*do* interact with other atoms, and that expectation is correct.

The third row, beginning with sodium and ending with argon, follows the same pattern. It's all about electrons filling the lowest possible energy levels in an atom.[15] The fourth row, which begins with potassium and ends with krypton,

---

[14] The two electrons in helium make up a filled shell, so they don't interact readily with electrons in other atoms. I mention this because we commonly think of valence electrons as interactive. Not so with the noble gases.

[15] We're focusing on what happens to electrons in neutral atoms, but don't forget that as you go from element to element, the identifying change in the atoms is the change in the number of protons.

all of a sudden has atoms in the gap. Why? Because at this point we begin filling the energy level 3*d*. The 3*d* energy level can hold ten electrons, which is why there are ten atoms filling the gap until you get to gallium. At that point, adding protons and their corresponding electrons involves the 4*p* energy level. Because each *p* level holds six electrons, we traverse six atoms in getting to the atom of krypton. Krypton has a filled 1*s* level, a filled 2*s* level, a filled 2*p* level, a filled 3*s* level, a filled 3*p* level, a filled 4*s* level, a filled 3*d* level, and a filled 4*p* level.

I should say something about the two rows at the very bottom that are separated from the rest of the Periodic Table. These represent a progression of elements in which electrons are filling *f* energy levels. If we included them in the main portion of the Periodic Table, there would be another huge gap that would stretch the table from side to side; then we wouldn't be able to fit a readable Periodic Table on a page. It's sort of like how Alaska and Hawaii are portrayed on a map of the United States. They're not shown in their actual locations because if they were, the rest of the map would not be readable on a normal-size page.

All atoms in a particular column of the Periodic Table have similar properties because they have the same number of valence electrons (electrons in the outermost *s* and *p* energy levels). This general rule is the same for all columns of the Periodic Table. Magnesium, calcium, and strontium have similar properties—they're all shiny, soft metals that react with water to produce hydrogen. Copper, silver, and gold have similar properties—they're also shiny, soft metals that are excellent conductors of heat and electricity. Fluorine, chlorine, bromine, and iodine have similar properties—all react strongly with other elements and exist as diatomic molecules in their pure states. For elements in columns 3 through 12, the number of valence electrons (the *s* and *p* electrons) is not changing because you're adding *d* electrons. Even so, the atoms in these rows have similar properties because they have similar electron arrangements.

Bonding and such

We've all known since maybe third or fourth grade that atoms get together and form larger things called *molecules*. This is what most of chemistry is about. When atoms hold onto one another, we say they form **bonds**. You can understand bond formation by focusing on atoms' desire to have filled outer shells and considering the energy involved.

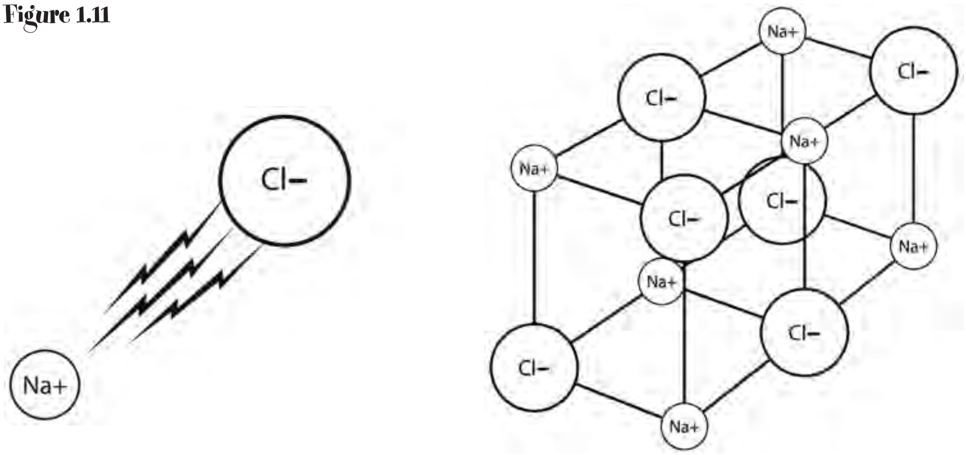
One kind of bond is the **ionic bond**.[16] This is where one atom steals an electron from another atom. This makes one a positive ion and the other a nega-

---

[16] There are chemists who prefer not to consider ionic bonds as true chemical bonds, but because opinions are mixed on that issue and ionic bonds seem to fall into the category of chemical bonds for me, we'll stick with that convention.

tive ion, with an electric attraction between the two (opposite charges attract). The result is that the two atoms bond together. Of course, ionic bonds usually involve many more than just two atoms. They usually involve many atoms linked together, as in the salt lattice represented in Figure 1.11.

Figure 1.11

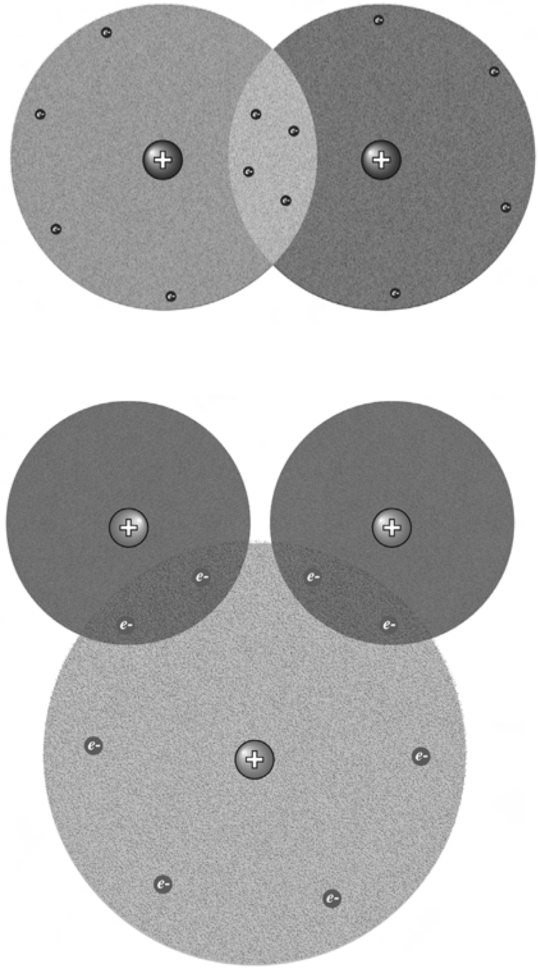


When atoms form ionic bonds, the new bond has a lower energy than the two atoms do separately. This follows the general principle that systems tend toward lower energies. You can predict which atoms will get together and form ionic bonds. For example, atoms in the first column of the periodic table have one valence electron and would love to get another one to fill their *s* shells. Atoms in the next to last column of the Periodic Table need just one electron to fill their outer shells. Remember that filled shells are energetically favorable. So anyway, this means that when you bring together atoms from the first column and atoms from the next to last column, you are likely to get an ionic bond. By the way, there's a number known as an atom's **electronegativity** that lets you know how readily an atom will grab or give up an electron. For details on electronegativity, check a more thorough resource than this chapter, like maybe the first chemistry book in this series.

SCILINKS®
THE WORLD'S A CLICK AWAY
Topic: Electronegativity
Go to: *www.scilinks.org*
Code: MCB006

The second kind of bond, which has two subclassifications, is a **covalent bond**. Sometimes when atoms get together, it's possible for them to share one or more electrons rather than steal an electron from each other, and that's a covalent bond. By sharing electrons, each atom can satisfy its need for a complete outer shell even though neither atom has a strong enough attraction to or repulsion from electrons for an electron to jump from one atom to another. This sharing of electrons, which forms a bond between the atoms, gives each atom a configuration almost as good as a filled outer shell and is thus more favorable energetically than the atoms staying apart.[17] A few examples of covalent bonds are two hydrogen atoms getting together to share their single electrons and form a hydrogen molecule (H_2), two oxygen atoms getting together to share four electrons (two each) and form an oxygen molecule (O_2), and two hydrogen atoms getting together with one oxygen atom to share a total of four electrons and form a water molecule (H_2O). Figure 1.12 shows the electron-sharing for oxygen and water. I should, of course, be shot for providing the drawing in Figure 1.12. I already told you that electrons don't occupy specific positions in atoms, but the drawing seems to imply that. All I can ask is that you use the drawing as an indication of shared electrons (the ones in the overlapping region) without making the leap that this is what atoms in a covalent bond *actually* look like. Also, keep in mind that the drawing only shows the valence electrons; there are others in each atom not shown.

Figure 1.12



Now for the subclassifications. Sometimes atoms share electrons even though one atom likes electrons (attracts them) more than the other atom does. In this case, you still have the sharing of electrons and a covalent bond, but the electrons spend more time around one of the atoms than they do around the other. Such a bond is called a **polar covalent bond**, with the word *polar* meaning that one part of the molecule is more positive and the other side is more negative. Covalent bonds in which the atoms share the electrons more or less equally are called **nonpolar covalent bonds**. O_2 is a nonpolar covalent bond (the shared electrons hang around each oxygen molecule for about equal amounts of time), and

---

[17] Remember that filled shells generally represent lower energies than unfilled shells, and systems tend toward lower energies.

H_2O is a polar covalent bond (the shared electrons spend more time around the one oxygen atom than they do around the two hydrogen atoms).

Institute for the Study of
Polar Covalent Bonds
North Campus

Institute for the Study of
Nonpolar Covalent Bonds
Bali Campus

Finally, metals have their own unique kind of bond. Each atom in a metal shares one or more electrons with *all* of the other atoms of that same piece of metal in what's called a **sea of electrons**. So, metals have a whole bunch of free-roaming electrons that can move all over the place. Again, this sharing of the electrons by all of the metal atoms is more favorable energetically than some other configuration.

When chemical structures collide

Chemistry is more than atoms getting together to form molecules. Most often we think of **chemical reactions**, in which one or more chemical compounds get together, atoms switch places, and something observable happens. For example, when you mix baking soda and vinegar, you end up with a new substance called sodium acetate, along with carbon dioxide gas and water. We represent such reactions with a **chemical equation**, which shows how a certain amount of one substance reacts with a certain amount of some other substance(s) to produce a certain amount of new substance(s). For example, our baking soda and vinegar reaction can be represented, first in words and then in chemical formulas, as

vinegar + sodium bicarbonate → sodium acetate + carbon dioxide + water

$$HC_2H_3O_2\ (aq) + NaHCO_3\ (s) \rightarrow NaC_2H_3O_2\ (s) + CO_2\ (g) + H_2O\ (l)$$

In this reaction, we refer to the molecules on the left as **reactants** and the molecules on the right as **products**. The equation shows what happens to single

"No! We don't need collision insurance."

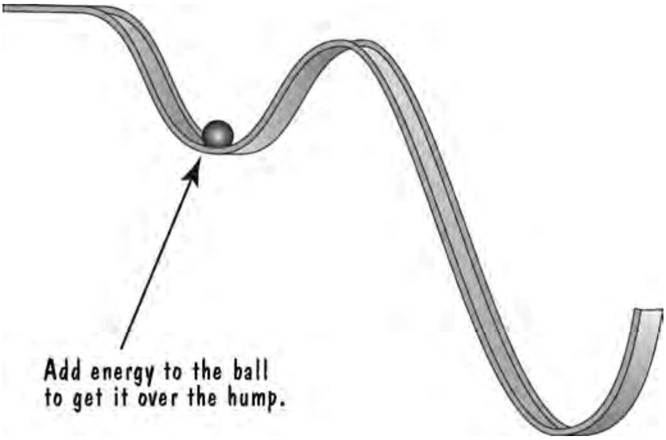
molecules of baking soda (sodium bicarbonate) and vinegar and thus is only *representative* of the thousands or millions of such reactions taking place at once.

If you count up the number of each kind of atom on the left and right sides of the previous equation, you'll find out that the numbers are equal. This is consistent with what is known as the **law of conservation of mass**, which basically states that atoms are not created or destroyed during a chemical reaction. You have to end up with the same number of each kind of atom as you started with. The equation for baking soda and vinegar reflects this fact, but sometimes when you write down the chemical formulas for the reactants and products, the number of atoms doesn't balance. Then you have to go through a process of "balancing" the equation—multiplying each chemical compound on the left and right by various numbers to get the number of atoms on the left and right to come out the same. This process can be easy or take quite a bit of trial and error. The more you practice, the better you get at it, which is how most things in life work.

Why, you might ask, do chemical compounds engage in a chemical reaction? Why don't they just keep to themselves when you mix them together? The answer lies again with energy. Sometimes when you mix two substances together, nothing happens. Sometimes a lot happens, but we huge humans can't see what's going on. Sometimes the substances spontaneously react, and sometimes they even explode. When they do react, it's partially because the products have a lower energy than the reactants. The new bonds that form in the products have an overall lower energy than the original bonds between the reactants. Putting energy aside, you can also understand reactions as taking place because different atoms have different affinities for other atoms. It's like a bunch of hormonal teenagers getting together at a dance. During the dance, some people find they like someone else better than the person they came with, so they switch over and form new partners.

Sometimes you have to add energy to get a chemical reaction to occur. For example, a stick of dynamite will sit around forever doing nothing until you light a fuse (add energy). It's like giving a ball a push (adding energy) over a hump in order for it to get to a lower place, as in Figure 1.13.

Figure 1.13

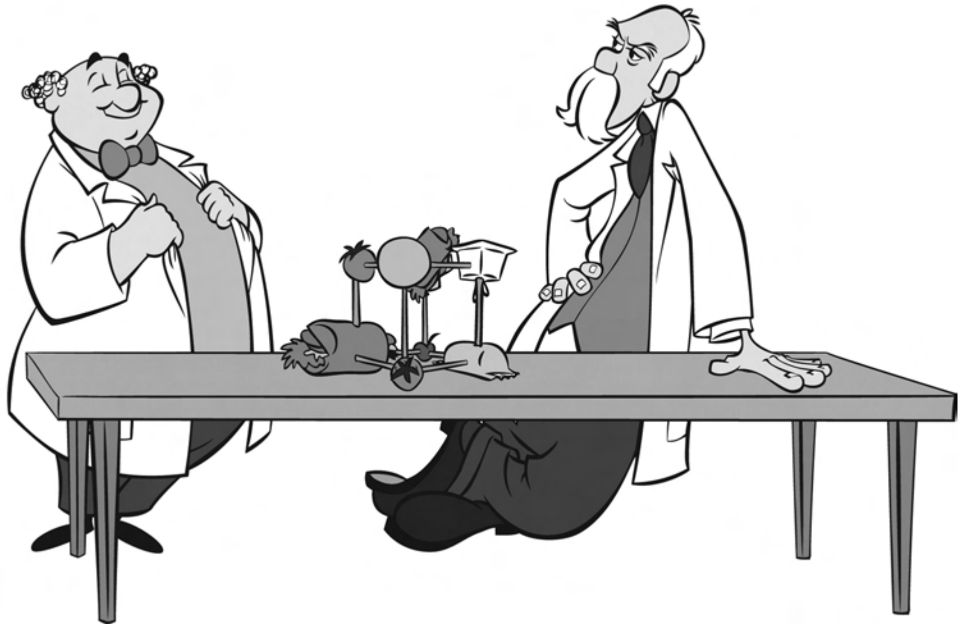


For example, many chemical reactions will occur once you heat up the reactants. Cooking an egg involves chemical reactions that take place only after you add heat. The energy you add to get a reaction going is known as the **activation energy**.

Organic stuff

In the first book, I briefly touched on **organic chemistry**, which is the study of chemical reactions involving carbon. Carbon is special in that its valence electrons are in just the right orientation to allow carbon to form many kinds of bonds with itself and other atoms. This leads to long, chainlike molecules (called **polymers**) that have many useful properties. Here I'm talking about things like polyester, vinyl, and different kinds of volatile substances, such as octane and methane (natural gas). Silicon, being similar to carbon (it's right below carbon in the Periodic Table), also fits into the study of organic chemistry.

Okay, that's my quick summary of the first chemistry book in the *Stop Faking It!* series, *Chemistry Basics*. If this wasn't a summary of concepts for you but rather an introduction, then you really should grab that first book. Doing the activities that precede the thorough explanations there will do much more for your understanding than this single chapter possibly could.



2
Chapter

Dynamic Atoms

This chapter is about states of matter and a section of chemistry called **thermodynamics**, which is the inspiration (or lack thereof) for the chapter title. Some of what is in this chapter might seem just a bit disconnected from the content in the rest of the book, but it's an important part of chemistry and helps introduce the equally important concepts of moles and molarity. The best way to introduce those concepts is to start with how gases behave. You also need a basic understanding of the states of matter. If you have read the *Stop Faking It!* book *Air, Water, and Weather*, you will recognize some of the activities in this chapter. Before skipping over those, you should know that the explanations that follow will not be identical to those in the other book.

"'Dynamic atoms'? Is that the best title you two could think of?"

In this chapter, I will assume you have a solid grasp of the following concepts: mass, speed, kinetic energy, and force. If my assumption isn't correct in your case, you might want to learn a bit about those concepts before continuing. The following text box contains a set of brief definitions for those concepts in case you just need a review.

mass—a numerical measure of an object's inertia. Inertia and mass give you an idea of how difficult it is to change the motion of an object. An object with a greater mass is harder to get moving and stop, and hits other objects harder for a given speed. Throughout this book, we will refer primarily to the masses of atoms and molecules and collections of atoms and molecules. The units of mass in the metric system are kilograms.

speed—how fast something is moving in miles per hour, meters per second, etc. The formal definition of speed is (distance traveled)/(time to travel that distance). The direction in which something is moving is irrelevant when calculating speed. It is usually relatively easy to calculate average speed over a given time period, but instantaneous speed (how fast something is moving at a particular instant) tends to be more difficult to determine.

velocity—speed with a specification of direction added. The magnitude of an object's velocity is just its instantaneous speed.

kinetic energy—the energy something has as a result of its motion. The formula for kinetic energy is $\frac{1}{2}mv^2$, where m is the mass of the object and v is the magnitude of the object's velocity. The units of kinetic energy are joules, the units for all forms of energy.

force—any push, pull, nudge, whack, or shove. The units of force in the metric system are newtons.

Things to do before you read the science stuff

To start things off, I'm going to have you think about experiences you already have with water changing from one state to another.[1] By **state,** I mean the categories of solid, liquid, and gas. First, think about water changing from its liquid state (which we call water!) to its solid state, otherwise known as ice. In changing from liquid water to solid ice, does the water molecule change at all? If you don't know the answer to that question, then leave an ice cube out to melt. What do you get when it melts? If you said water, give yourself a good grade. Next, either boil some water or think about the last time you boiled water. What happens? Yep, you get steam. Is the steam still composed of water molecules, or something different? When answering that question, think about what happens to steam when it comes in contact with a cool surface, such as when the steam from your shower comes in contact with the bathroom mirror or when droplets form on the outside of a glass of iced tea on a warm day.

The science stuff

Hopefully the activities described above help convince you that when a substance goes from solid to liquid to gas and back, the basic unit of our substance—the water molecule—doesn't change. If the molecules changed when you froze water, then you wouldn't expect to get water back when ice melts. Similarly, if steam,

---

[1] If you haven't had the experiences described, by all means do them now.

otherwise known as water vapor, were not composed of water molecules, then you wouldn't expect to get water back when the steam encounters a cool surface.

What happens with water also happens with many other substances. Changing states does not alter the chemical composition of the substance. For example, when you cool oxygen gas (O_2) to very low temperatures, it becomes liquid oxygen that is composed of O_2 molecules. The only difference between liquid and gaseous oxygen is how strongly the molecules are held together.

So, for all substances, they can be solids, liquids, or gases and still be the same basic substance. In a solid, the atoms or molecules are held together so strongly that their positions with respect to one another don't change much. The atoms or molecules in a solid move in a jiggling sort of way, as do all atoms and molecules, but they essentially stay put. The atoms and molecules in liquids move more, rolling around one another without losing touch with the liquid as a whole. The atoms and molecules in a gas are free to bounce around all over the place and don't latch onto one another at all. (When they do latch on, they become liquids!)

More things to do before you read more science stuff

For this section, get yourself to a computer and head to the following link: *http://intro.chem.okstate.edu/1314F00/Laboratory/GLP.htm.* Be sure to read the instructions before starting.

Note: If you don't have a computer handy, imagine a bunch of marbles bouncing around inside a container and following the rules I've described. Ball bearings in a closed petri dish also work well. If you're doing this with a classroom full of kids, you can have the kids pretend to be molecules in an enclosure. I describe such an activity in the *Stop Faking It!* books *Energy* and *Air, Water, and Weather.* Of course, the best way to visualize a gas is with the computer simulation, so do your best to use that.

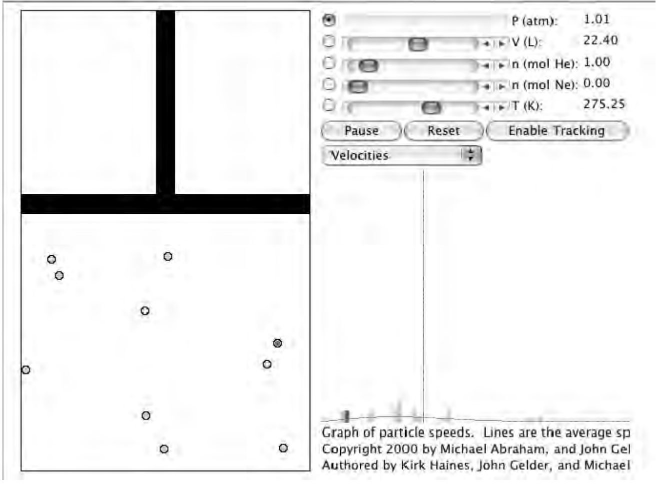
This is a simulation of how gas molecules behave. To be more precise, this is a simulation of how an *ideal* gas behaves. In an ideal gas, the individual atoms or molecules don't interact with one another unless they collide.[2] When the atoms or molecules collide, they bounce off one another as if they were billiard balls. They also bounce off the walls of the container just as billiard balls do.

---

[2] An ideal gas is a good first approximation for how all gases behave, and it is a good enough model for us right now. When considering how actual gases behave, you have to allow for the rotation of molecules and the motion of atoms within a molecule, in addition to the fact that real gases do interact in ways beyond just bouncing off one another. For most applications, the ideal gas model works pretty well.

During a collision, it's possible for an individual atom or molecule to speed up or slow down. An atom or molecule also maintains a constant speed in between collisions. Figure 2.1 shows a still drawing of the simulation.

Figure 2.1



● P (atm): 1.01
○ V (L): 22.40
○ n (mol He): 1.00
○ n (mol Ne): 0.00
○ T (K): 275.25

Graph of particle speeds. Lines are the average sp
Copyright 2000 by Michael Abraham, and John Gel
Authored by Kirk Haines, John Gelder, and Michael

© Michael R. Abraham and John Gelder. Used with permission.

Once you have the simulation running or are actively imagining gas molecules moving around in a closed container, I want you to notice a few things. First, click on the enable tracking button so you can follow easily the motion of a single molecule. Verify that this molecule follows the rules of an ideal gas that I already mentioned. Kinda fun watching the little fella bounce around, huh?

Second, notice that all of the quantities listed in the control bar region—pressure, volume, temperature, etc.—are ones that apply to the gas as a whole. In an actual gas, there are way too many atoms or molecules to track individually. Yes, the graphing region shows the velocities of individual molecules, but that's because there are a limited number of those in this simulation. In practice, you cannot keep track of individual velocities because you are dealing with billions of molecules. Instead, we deal with averages. For example, we register the pressure of the gas (the force per unit area exerted on the walls of the container) instead of the forces that individual molecules exert. We can't keep track of the forces exerted by individual molecules, but we can keep track of the average force they exert on a given area.

Pressure is defined as the force exerted on a given area divided by that area. If you push on your hand with the sharp end of a nail, the area over which that force is exerted is small, and the force per unit area, the pressure, is large—it hurts! If you push with the same force using the flat end of the nail, however, that force is spread out over a larger area, and the resulting pressure is low—it doesn't hurt as much. When considering a gas, you can think about the pressure the gas exerts as being a measure of how hard, overall, the gas pushes on its surroundings and the molecules push on one another.

Third, I want you to take a look at the units of the quantities shown in the control bar. The pressure is measured in the unit **atm**. This is not a reference to quick cash, but rather an abbreviation for *atmospheres*. One atmosphere is a pressure roughly equal to the air pressure at sea level. Volume is measured in liters, a unit with which you should be familiar. The third and fourth control bars indicate the number of atoms of helium and neon that are present. The unit is mol, which stands for the word *mole*. For now, just think of this number as an indicator—not an exact count—of the number of atoms, in either the simulation or the real gas the simulation represents. For example, the default value of the number of moles of helium is 1.0. Clearly, there's more than one atom of helium in the simulation. Later on, you'll find out how many atoms of a real gas this 1.0 represents (a *lot*!). The temperature is measured in degrees Kelvin, or K. Water freezes at 273.16 degrees Kelvin, which is 0 degrees Celsius or about 32 degrees Fahrenheit.

Finally, make sure you understand how the sliding controls work. If you click on what the authors call a radio button (the circle to the left of each control), then that quantity is free to change in response to changing the other quantities. All the other quantities remain fixed unless you move the slider associated with the quantity. So, in the default settings, you can change the volume, the number of each kind of atom, and the temperature. The pressure will change in response to the changes you make.

Okay, on to investigating the behavior of your ideal gas. You can learn quite a bit about gases by simply playing around with the various controls. I encourage you to do that, but I definitely want you to be sure to try the following and see what happens in each case:

- As you change the temperature, what happens to the velocities of the atoms in the gas? Note that the average velocity of each kind of gas (helium or neon) is indicated by a vertical line on the graph.
- What happens to the pressure exerted by the gas as you increase the temperature? What happens to the pressure exerted by the gas as you decrease the temperature?
- What happens to the pressure exerted by the gas as you increase or decrease the number of atoms?
- What happens to the pressure exerted by the gas as you increase or decrease the volume occupied by the gas? Remember that with the pressure radio button checked, the temperature and number of atoms remain constant as you change the volume.

Before moving on to other tasks, think about the results you have obtained so far, and see if they make sense to you in terms of what the atoms are doing.

You've been monitoring changes in pressure as you change other quantities. Should the pressure go up when you increase the number of atoms? Why? Should the pressure decrease when you increase the volume? Why? Go ahead and answer these and other *why* questions. As you do this, keep the following concept in mind: The pressure the gas exerts will be affected by how fast and how often the atoms hit the container walls.

More tasks to try with the simulation:

- Click on the radio button for volume. Find out what happens to the volume as you change the pressure, number of atoms, and temperature. As you do this, be sure to note which quantities are constant and which are changing. For example, with the volume radio button checked, you can find out what happens to the volume when you change the pressure while keeping the number of atoms and the temperature constant. To investigate this in the real world, you would have to set up your apparatus to ensure that these latter quantities remained constant. That can be difficult (how would you ensure that the temperature of the gas didn't change?), which is why a simulation like this one is a valuable tool.
- Click on the radio button for temperature. Find out what happens to the temperature as you change the pressure, volume, and number of atoms. Again, note which quantities are constant and which are changing.
- Click on the radio button for the number of atoms of helium and set the number of atoms of neon at zero. Find out what happens to the number of helium atoms when you change the pressure, volume, and temperature. As you're doing this, realize that you're answering the following question: What must happen to the number of atoms when I change one quantity while requiring that the other quantities remain constant? This would be quite difficult to do in the real world. You would have to have some mechanism for introducing and removing gas molecules from your container. I just want to make sure you don't get the notion that in a real situation atoms would just appear and disappear, which is what happens in the simulation.

As before, take a look at these results and try to make sense of them in terms of the motion and quantity of atoms in the gas. Then do one more thing: Arrange the simulation so that you have both helium and neon atoms present. At any given temperature, find out which atoms, on average, move faster and which on average move slower. Easy task, no?

More science stuff

Lots to explain here. I'll start with the first thing I asked you to do, which was to notice what happens to the speeds[3] of the atoms as you change the temperature. Higher temperatures mean higher speeds, and lower temperatures mean lower speeds. Given that relationship, you might think that the temperature of the gas is a measure of the average speed of the particles in the gas. Not quite true. Imagine you have a thermometer stuck inside the gas to measure the gas's temperature. The hotter the liquid in the thermometer gets, the more it expands and moves up a column on which different temperatures are stamped. The faster the gas particles are going when they hit the thermometer, the more energy they give to the thermometer, and the more the liquid expands. But speed isn't everything. The more massive the particles are, the more energy they give the thermometer in a collision. It's just like getting hit with a pebble going 2 meters per second compared to getting hit with a large rock going 2 meters per second. The rock imparts more energy and hurts more. If we wanted to attain the same "hurt" with the pebble and the large rock, the pebble would have to be moving much faster than the rock.

So our definition of what temperature is should somehow contain both the speed of the particles and their mass, and it does. The temperature of a gas is directly related to the average **kinetic energy** of the gas particles. An object's kinetic energy, or motion energy, is given by the formula $\frac{1}{2}mv^2$, where m is the object's mass and v is the magnitude of the object's velocity. So the greater the mass, the greater the temperature, and the greater the speed, the greater the temperature. Now think back to the last thing I had you do in the previous section. I had you notice that at a given temperature, the neon particles moved slower than the helium particles. This makes sense in terms of our definition of temperature as a measure of the average kinetic energy of the particles. The helium and neon particles, as a whole, have the same temperature and, thus, the same average kinetic energy. Neon is more massive than helium, though; thus, the neon particles will have a smaller velocity because their mass is larger. See Figure 2.2 (p. 30).

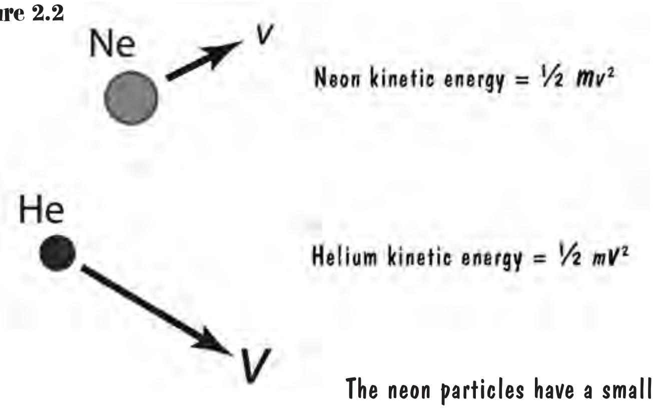
SCILINKS®
THE WORLD'S A CLICK AWAY

Topic: Kinetic Energy
Go to: *www.scilinks.org*
Code: MCB007

---

[3] There is a difference between speed and velocity. Speed is only concerned with magnitude (how big the quantity is), and velocity includes both magnitude and direction. The authors of the simulation use the word *velocity,* but because we're not recording direction, the correct term to use would be *speed*. Scientists often use the two terms interchangeably when it is clear that we are only considering the magnitude of the velocities, and that's what's happening here. In the text, I will refer to speeds; be aware that the corresponding term in the simulation is *velocities*.

Figure 2.2



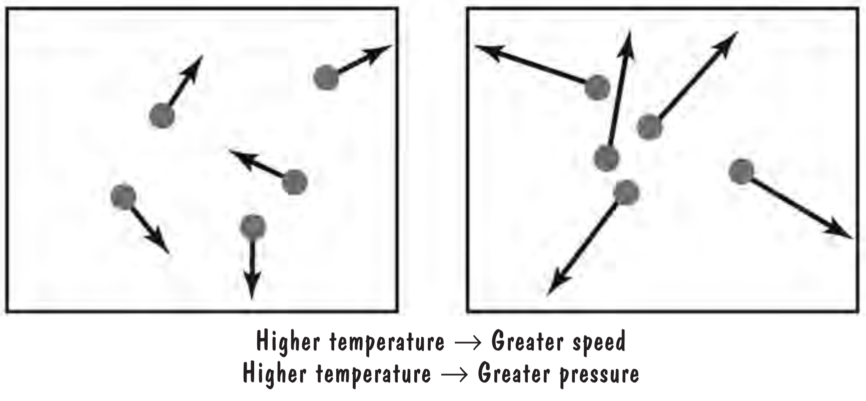
The neon particles have a smaller velocity (indicated by a small v), but they have the same average kinetic energy as the helium particles because they have a larger mass.

On to the other things you did with the simulation. Here are the results you no doubt found.[4]

- With all other quantities held constant, increasing the temperature of the gas increases the pressure the gas exerts on the container. This should make sense to you. Increasing temperature increases the speed of the gas particles. This increased speed makes the particles hit the wall harder, thus increasing the pressure. Conversely, decreasing the temperature decreases the pressure the gas exerts. Take a look at Figure 2.3.

Figure 2.3

The particles on the right move faster and hit the container harder.



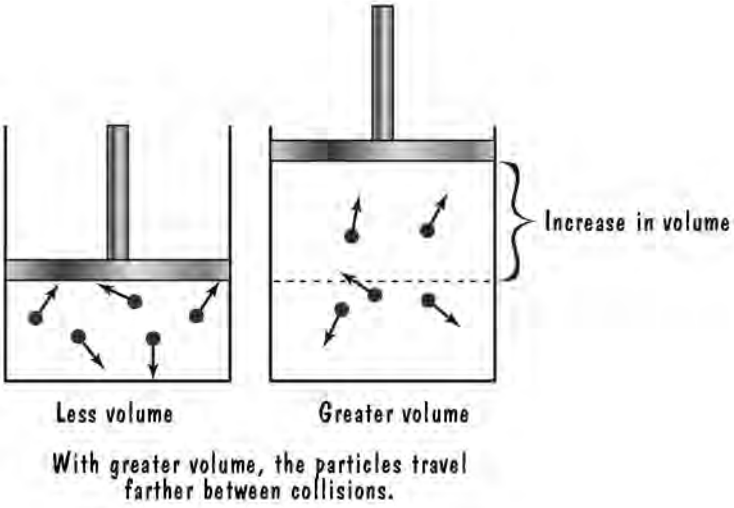
Higher temperature → Greater speed
Higher temperature → Greater pressure

---

[4] As with all of my books, it's possible for you to find out what should have happened in the activities I ask you to do by reading the section that follows. You will have a better grasp of concepts, though, if you actually do the activities when I tell you to do them. In other words, you understand better when you do activities rather than just read about them.

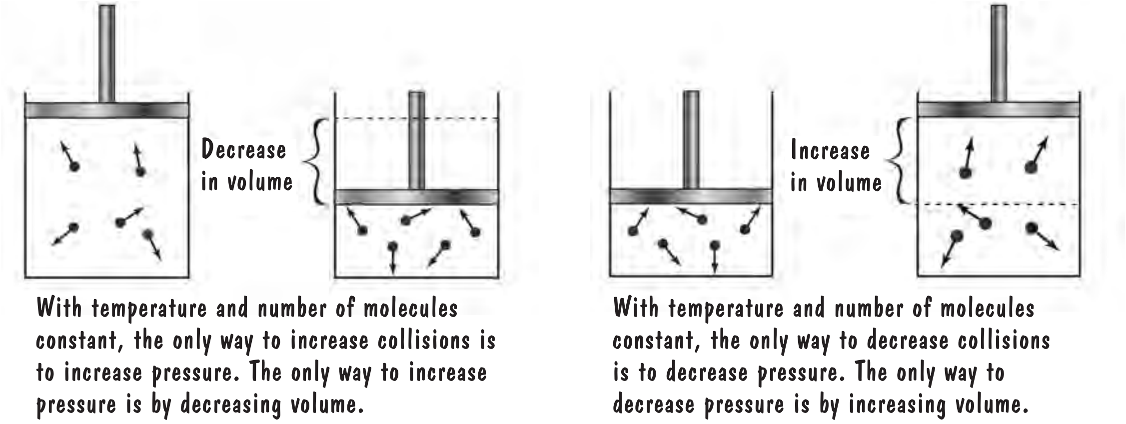
- With all other quantities held constant, increasing the number of gas particles increases the pressure. These added particles are at the same temperature as the other particles, so they don't hit the container any harder. Because there are more particles, though, they hit the walls of the container more often. More collisions with the container mean a higher pressure. Removing gas particles decreases the pressure because there are fewer collisions with the walls of the container.
- With all other quantities held constant, increasing the volume of the gas decreases the pressure. Because there is now more distance between the walls, the gas particles now hit the walls of the container less often. Fewer collisions mean a lower pressure. Conversely, decreasing the volume increases the pressure, because now the particles hit the walls more often. See Figure 2.4.

Figure 2.4



- With all other quantities held constant, increasing the pressure decreases the volume, and decreasing the pressure increases the volume. You have to think a bit to interpret this result. Remember that when you allow the volume to change and then change the pressure, you are requiring that the other quantities stay where they are. You don't change the number of particles or the temperature. With those restrictions, the only way to increase the pressure is by decreasing the volume, and the only way to decrease the pressure is by increasing the volume. See Figure 2.5 (p. 32). Basically, the simulation is doing what it can, given the restrictions you give it. When dealing with an actual gas, it would be really difficult to increase the pressure without changing the temperature.

Figure 2.5



With temperature and number of molecules constant, the only way to increase collisions is to increase pressure. The only way to increase pressure is by decreasing volume.

With temperature and number of molecules constant, the only way to decrease collisions is to decrease pressure. The only way to decrease pressure is by increasing volume.

I haven't addressed everything I had you do in the previous section because I think you can figure the rest out for yourself. Just remember how the different quantities are related, that pressure depends on how often and how fast the particles hit the container, and that temperature is a measure of the average kinetic energy of the gas particles.

Even without addressing everything you did, we have enough information to talk about a few laws that scientists came up with many years ago. Chemists who worked on the behavior of gases in the old days didn't have computer simulations to study, primarily because they didn't have computers. They did, however, perform experiments using real gases under controlled conditions.

In the 1700s, Joseph-Louis Gay-Lussac discovered that there was a direct relationship between the pressure and temperature of a gas. He expressed it as

$$\frac{\text{pressure}}{\text{temperature}} = \text{constant}$$

or

$$\frac{P}{T} = \text{constant}$$

This fits with your experience with the simulation. If the pressure divided by the temperature is a constant number, then when you increase the pressure, the temperature also must increase to have P/T equal to the same number. Remember that this is only true if you keep the number of molecules and the volume constant, which is what Gay-Lussac did in his experiments.

Earlier than that, in 1660, Robert Boyle discovered a relationship between the pressure and volume of a gas that's true when holding the temperature and number of molecules constant, which is

$$\text{(pressure)(volume)} = \text{constant}$$

or

$$PV = \text{constant}$$

This is an *inverse relationship,* meaning that as pressure increases, the volume decreases, and as pressure decreases, volume increases. These relationships are true when the temperature and number of molecules are held constant. Again, this fits with what you observed in the simulation.

Still another scientist, Jacques Charles, discovered in 1787 a direct relationship between volume and temperature (number of molecules and pressure held constant), which is

$$\frac{\text{volume}}{\text{temperature}} = \text{constant}$$

or

$$\frac{V}{T} = \text{constant}$$

This is just like the relationship between pressure and temperature. As the pressure increases, the volume must also increase so the ratio V/T remains a constant number.

Finally, Amedeo Avogadro[5] proposed in the 1800s that the number of molecules in a gas was directly proportional to the volume of the gas. This relationship is expressed as

$$\frac{\text{number of molecules}}{\text{volume}} = \text{constant}$$

or

$$\frac{N}{V} = \text{constant}$$

Once again, this fits with what you observed in the simulation. As the number of molecules increases or decreases, the volume of the gas increases or decreases accordingly.

The above relationships are labeled according to the person who discovered them and are thus known as Gay-Lussac's law, Boyle's law, Charles' law, and

---

[5] As we all know, this famous scientist was immortalized in the song "Rock Me, Amedeo." As you should know if you've read any of my other books, this is just a bad joke with an obscure reference.

Avogadro's principle. As you go over these relationships, it's really important to remember that each law assumes that all other variables associated with the gas are held constant. Now, I suppose for historical reasons all chemistry books make a point of naming all of these laws separately. Fortunately for you, we can tie all of them together in one law known as the **ideal gas law**. Here it is:

(pressure)(volume) = (number of moles)(a constant)(temperature)

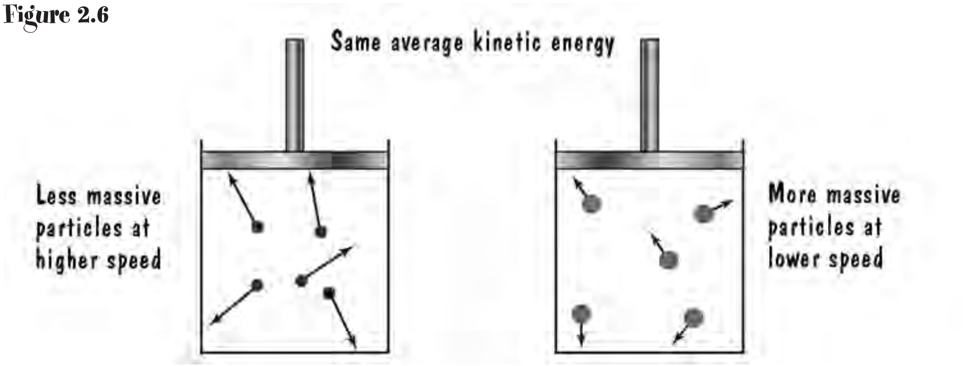
or

$$PV = nRT$$

The letter R is the constant number, known as the universal gas constant or the ideal gas constant. The value of this constant is $0.0821\ \frac{L-atm}{mol-K}$. You read the units on this as *liter atmospheres per moles degrees Kelvin.* Liters are units of volume, atmospheres are units of pressure, moles are units related to the number of atoms or molecules, and degrees Kelvin are units of temperature. As with most of my other books, I'm not going to focus on units very much, but I want to be complete and use the proper units. One important thing to remember if you study thermodynamics further is that the ideal gas constant can be expressed in a variety of units, as can other constants in chemistry and physics. If you ever need to use this form of the equation to solve a problem or do an example, be sure to write out all the units and be sure that they all "fit together" before you do the math.

I'm going to give you an example of how to use the ideal gas law in the Applications section. To end this section, though, I want to expand on Avogadro's principle. The basis of the principle is the hypothesis that gases that have the same pressure, volume, and temperature have the same number of molecules. But how could different gases, with different masses for their atoms or molecules, have the same number of atoms or molecules at the same temperature? I mean, wouldn't the heavier gases at the same temperature exert a greater pressure because of the added mass? And wouldn't this mean that there should be *fewer* of the heavier gases in order to have the same pressure, volume, and temperature? If this is what's going through your mind, then possibly you're thinking of temperature in terms of speed only and not as a measure of kinetic energy, which includes the mass of the particles. Think back to the simulation, where more massive particles at the same temperature as less massive particles moved more slowly. So, the same number of massive and not-so-massive particles could be at the same temperature (a measure of kinetic energy and not just speed) and still exert the same pressure at a given volume. See Figure 2.6.

Figure 2.6



Even more things to do before you read even more science stuff

Gather together 30 pennies or other small objects, such as marbles. Remembering your higher math, you know that there are 12 objects in a dozen, such as for eggs and donuts. How many dozen pennies do you have? Easy.

Now I'm going to define a new quantity called a *grik*. A grik of something contains three objects. So, if you have 12 eggs, that's 4 griks of eggs. How many griks of pennies do you have? Now let's use a new quantity called a *foom*. A foom of something contains 8 objects. How many fooms of pennies do you have?

Even more science stuff

What I had you do in the previous section probably seemed trivial. If so, that's good. It's common for us to specify a certain quantity of things with a special name, such as a dozen, gross, or bushel. If you had 30 pennies, you had 2.5 dozen pennies, 10 griks of pennies, and 3.75 fooms of pennies. There's a quantity often used in chemistry to specify how many atoms or molecules of a substance you have, and it's called a **mole**. One mole contains 6.02214×10^{23} atoms or molecules, and this number is known as **Avogadro's number**. One mole of hydrogen molecules contains 6.02214×10^{23} hydrogen molecules. One mole of xenon atoms contains 6.02214×10^{23} xenon atoms. For that matter, one mole of pennies contains 6.02214×10^{23} pennies, but there aren't that many pennies in existence.

I already explained the reasoning that makes Avogadro's principle a good hypothesis, and of course it's been verified many times that gases at a given pressure, volume, and temperature have the same number

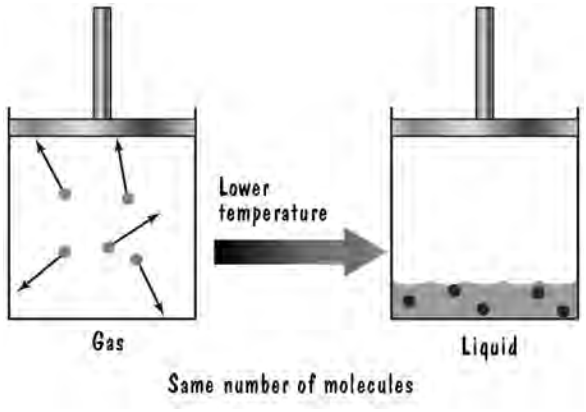
SCILINKS
THE WORLD'S A CLICK AWAY

Topic: Avogadro's Constant
Go to: *www.scilinks.org*
Code: MCB008

of atoms or molecules. We can define one mole as the number of molecules in a gas at an agreed-upon pressure, temperature, and volume, known as Standard Temperature and Pressure (STP). The values of these quantities at STP are a pressure of one atmosphere (roughly the atmospheric pressure at sea level), a temperature of 273.16 degrees Kelvin (the freezing temperature of water at a pressure of one atmosphere), and a volume of 22.4 liters. Turns out that Avogadro never knew exactly how many molecules were in a gas at various conditions, so the definition of a mole came after his demise. Poor soul didn't even know the number that has his name.

You might be wondering how in the world chemists chose the values for STP. Well, it has to do with the fact that measuring things in moles goes beyond talking about gases. It also applies to solids and liquids. Let's say you have a mole of helium gas confined in a chamber. Now cool this gas down to very low temperatures so that you now have liquid helium (yes, that's entirely possible, and it's common for laboratories to use liquid helium to cool other things). If you had a mole of helium gas, then you now have a mole of liquid helium. We know this because of the reasoning I gave you in the very first explanation in this chapter: When you change states from solid to liquid to gas and back, the number of atoms or molecules you have remains the same. Check out Figure 2.7. Although the whole idea of measuring quantities in moles originated with gases, it applies to all states of matter and is a convenient way to figure out how much of a substance you have.

Figure 2.7



We know that a mole (6.02214×10^{23} atoms or molecules) of any gas takes up 22.4 liters at standard temperature and pressure. There's another, often more useful, way to define a mole using the element carbon. The carbon atom has an atomic mass of exactly 12 if you ignore the other isotopes of carbon.[6] In 12 grams of carbon-12 there is exactly one mole of carbon-12 atoms, which defines what a mole is. Because all atomic masses in the Periodic Table are based on the

---

[6] The atomic mass is that number at the bottom of an element's box in the Periodic Table. The number tells you how many protons plus neutrons are in the atom. Because the mass of electrons is so small compared to the masses of protons and neutrons, we ignore the electrons' mass when determining the mass of an atom. Isotopes of elements are atoms that contain different numbers of neutrons. For example, carbon comes in the forms carbon-12, carbon-13, and carbon-14. Carbon-14 has two more neutrons in its nucleus than carbon-12.

atomic mass of carbon-12, you can similarly figure out how many moles of other substances you have by determining their masses. Let's look at another element, sulfur. Sulfur's atomic mass is 32.06 (not a whole number because of the existence of isotopes). If you have 32.06 grams of sulfur (this is called a **gram molecular mass** of sulfur), then you have exactly one mole of sulfur, or 6.02214×10^{23} sulfur atoms. If you have 96.18 grams of sulfur, you have three moles of sulfur atoms. Finally, let's consider a molecule such as water. The molecular mass of water is 18.01528 (add the atomic mass of oxygen to two atomic masses of hydrogen). If you have 18.01528 grams of water, you have a gram molecular mass of water and therefore have one mole of water molecules. If you have 18 grams of water, that's a bit more than half an ounce of water, so if you pour out slightly more than half an ounce of water in a dish, the dish contains slightly more than 6.02214×10^{23} molecules of water.

When using moles as a measure of how much stuff you have, you sometimes have to be careful. For example, I just told you that 18.01528 grams of water contains one mole of water molecules, but that amount of water does *not* contain one mole of hydrogen atoms. There are two hydrogen atoms in each water molecule, so a mole of water contains two moles of hydrogen atoms.

To end this chapter, I'm going to introduce a chemistry measurement that is based on moles. If you look at the concentrations of various liquids, you will find the terms **molar** and **molarity**, as in a 0.5 molar solution of sodium hydroxide.[7] Before explaining this term, I should point out that a *solution* is something dissolved in a liquid (often water), and if you specify the *concentration* of a solution, you are saying how much of a particular something you have dissolved in a liquid. Anyway, here is the definition of molarity:

$$\text{molarity} = \frac{\text{number of moles of a substance}}{\text{number of liters of solution}}$$

Molarity is specified with a capital M.

I lied. I have one more thing to tell you. There are two commonly used forms of the ideal gas law. One is the form I already introduced, namely $PV = nRT$. The second, used more by physicists than chemists, is below.

$$PV = NkT$$

---

[7] I'm only using this as an example, not asking you to actually mess around with sodium hydroxide, but for safety's sake I should go ahead and mention that this chemical can be pretty caustic, especially in the solid pellet form. Take all necessary precautions (gloves, goggles) when handling this stuff.

The only difference is the substitution of Nk for nR. The upper-case N represents the number of molecules of a gas, while the lower-case n represents the number of moles of a gas. k is known as **Boltzmann's constant** and has a value of 1.381×10^{-23} J/K, with the units reading "joules per degree Kelvin." As with the universal gas constant, Boltzmann's constant has different values depending on the units you use.

Chapter summary

- When the state of a substance—solid, liquid, or gas—changes, the identity of its component molecules does not change
- An ideal gas is one in which we treat the molecules as independent things that act like billiard balls that do not rotate or vibrate. In collisions these molecules can exchange energy, but no energy is lost to the environment.
- The temperature of a gas is directly related to the average kinetic energy of the molecules in the gas. The faster the molecules move, and the more massive they are, the greater the temperature.
- The pressure of a gas is how much force per unit area the molecules exert.
- With volume held constant, the pressure of a gas is directly proportional to the temperature of the gas. Higher temperatures mean higher pressure and lower temperatures mean lower pressure.
- With temperature held constant, the pressure of a gas is inversely proportional to the volume of the gas. Smaller volumes mean higher pressures and larger volumes mean lower pressures.
- With pressure held constant, volume is directly proportional to temperature. Higher temperatures mean larger volumes and lower temperatures mean smaller volumes.
- With volume and temperature held constant, pressure is directly proportional to the number of molecules in a gas. More molecules means higher pressure and fewer molecules means lower pressure.

Applications

1. In the ideal gas simulation I had you mess around with, the lowest temperature possible is 25 degrees Kelvin. Why doesn't the simulation go all the way to 0 degrees Kelvin? The answer is that 0 degrees Kelvin is theoretically impossible to attain. At 0 degrees Kelvin, otherwise known as **absolute zero**, all motion would stop. If all motion in a substance stopped, though, then we would know *exactly* where the molecules are and what they're doing (noth-

ing!). This is a violation of a principle I'm going to introduce in Chapter 3. So, although scientists are able to bring substances to temperatures extremely close to absolute zero, it's theoretically impossible to actually reach absolute zero. Of course, the theory (quantum mechanics) that prevents that might one day be shown to be incorrect, but until that time we're sticking with zero chance of absolute zero temperature for anything.

2. I told you that you can use the ideal gas law instead of remembering all of the separate laws that apply to gases, so I guess I'd better give you an example of that.[8] You can use either form of the ideal gas law, so I'll use the one that's more common in chemistry textbooks, which is $PV = nRT$. Let's suppose you have a gas that's at a temperature of 293 degrees Kelvin and you raise the temperature to 323 degrees Kelvin. You keep the volume constant and the number of moles (also the number of molecules) constant. How does this rise in temperature affect the pressure the gas exerts on its surroundings? Let's call the lower temperature of the gas Situation 1 and the higher temperature of the gas Situation 2. The ideal gas law holds true no matter what temperature and pressure we use, so we can write the following:

$$P_1V_1 = n_1RT_1 \text{ and } P_2V_2 = n_2RT_2$$

The subscripts refer to the different situations. There is no subscript on the R because it's a constant that has the same value in all situations. What I'm going to do is divide the first equation by the second equation,[9] as in

$$\frac{P_1V_1}{P_2V_2} = \frac{n_1RT_1}{n_2RT_2}$$

In our situation, we know that the volume and the number of moles are constant, meaning that $V_1 = V_2$ and $n_1 = n_2$. And of course, $R = R$. Therefore we can cancel the terms that are equal, as in

$$\frac{P_1\cancel{V_1}}{P_2\cancel{V_2}} = \frac{\cancel{n_1}\cancel{R}T_1}{\cancel{n_2}\cancel{R}T_2}$$

That leaves us with

$$\frac{P_1}{P_2} = \frac{T_1}{T_2}$$

---

[8] Remember that the ideal gas law only applies to an "ideal" gas rather than real gases. It's a good approximation, however, for what will happen with most gases.

[9] Dividing one equation by another equation always results in a new equation that's valid. If this step, or any other step I do here, makes you grind your teeth, it wouldn't be a bad idea to brush up on your basic algebra. I happen to know of a great book for this purpose, the *Stop Faking It!* book on math.

We know the initial and final temperatures and let's assume we know the initial pressure; therefore, all we have to do is solve the above equation for P_2, our final pressure. You can do this algebraic manipulation with a shortcut or two, or you can take it slow so you understand the process. I'll describe the process and then give the result. What you need to do is multiply both sides of this equation by P_2 and then multiply the resulting equation by $\frac{T_1}{T_2}$. That gives you

$$P_2 = \frac{T_2}{T_1} P_1 = \frac{(323\text{ K})}{(293\text{ K})} P_1 = 1.1 P_1$$

So, by increasing the temperature slightly, you have increased the pressure by a factor of 1.1. As with all science calculations, you should stop and make sure the result makes sense. Should increasing the temperature while holding other variables constant increase the pressure? Sure. If I had given you an exact value for the initial pressure (pressure is measured in a variety of units, common ones being Pascals or atmospheres), you could compute an exact value for the final pressure.

3. It might seem that which ideal gas law you use, $PV = NkT$ or $PV = nRT$, is just a matter of taste. There is an important distinction, though. In one case, you are relating the actual number of molecules in a gas to the easily measurable quantities of pressure, volume, and temperature. In the second case, you are relating the number of moles of a gas (an easily measurable quantity—just determine the mass of the gas and use the gram molecular mass of the gas to figure out the number of moles) to the quantities of pressure, volume, and temperature. So the two forms of the gas law make a connection between what we can observe directly, known as extrinsic properties of the gas, and what we cannot observe directly (the actual number of molecules), known as the intrinsic properties of the gas. Not terribly important for everyday calculations, but philosophically important in that we are connecting the observable with the unobservable.

4. Suppose you have 1,000 grams (one kilogram) of methane gas. How many moles of gas do you have, and how many molecules of gas do you have? To find out how many moles of gas you have, you need the gram molecular mass of methane. To figure that out, you need the chemical formula of methane, which is CH_4. One carbon atom has an atomic mass of 12 (I'm going to round off the atomic masses here to simplify things) and one hydrogen atom has an atomic mass of 1. Therefore, CH_4 has a molecular mass of 16 (one carbon and four hydrogens). This means that 16 grams of methane contain one mole of methane molecules. We have 1,000 grams of methane gas, so we have $\frac{1000\text{ grams}}{16\text{ grams per mole}} = 62.5$ moles.

When doing math in a science calculation, people often rely on memorized procedures that don't necessarily make sense to them. For example, why did I just divide 1,000 grams by the number of grams per mole to get the number of moles? It's comparable to figuring out how many cookies you have if each cookie weighs 2 pounds (big cookies!) and you have 40 pounds of cookies. You divide 40 pounds by 2 pounds per cookie to get 20 cookies.

Once we know how many moles of methane we have, it's easy to figure out how many molecules there are. Each mole contains 6.02×10^{23} molecules, so there are $(62.5 \text{ moles})(6.02 \times 10^{23} \text{ molecules per mole}) = 376 \times 10^{23}$ or 3.76×10^{25} molecules.

5. The concept of moles is extremely useful for determining how much of one substance you need to complete a chemical reaction with another substance. For example, the balanced equation for the reaction of zinc with hydrochloric acid is given below. *Note:* I'm not suggesting you actually try this reaction. It's a bit violent, and it's always a risk doing such a reaction with a strong acid.

$$\text{zinc + hydrochloric acid} \rightarrow \text{zinc chloride + hydrogen gas}$$

$$Zn\ (s) + 2HCl\ (aq) \rightarrow ZnCl_2\ (s) + H_2\ (g)$$

According to this balanced equation, you need two molecules of hydrochloric acid for every atom of zinc. Of course, we can't measure out individual atoms, so it's more convenient to think in terms of moles. The equation says that for every mole of zinc, you need two moles of hydrochloric acid. Now, let's suppose your sample of zinc has a mass of 130.76 grams, and you have 148.00 grams of HCl.[10] At first glance, it looks like there's not enough HCl to react with and completely use up all of the zinc because they have to be in the ratio of 2 to 1. But we shouldn't compare masses; we should compare the number of *moles*. We need two moles of hydrochloric acid for each mole of zinc, so we better figure out how many moles of each substance we have. The atomic mass of zinc is 65.38 (look it up in the Periodic Table), so if you have 65.38 grams of zinc, you have one mole of zinc atoms (refer back to the discussion of gram molecular mass on page 37 if that doesn't make sense). We have twice that much (130.76 grams), so we have two moles of zinc. Hydrochloric acid is composed of two atoms, so to find the *molecular* mass of HCl, we need to add the atomic masses of hydrogen and chlorine. Again referring to the Periodic Table, we find that the atomic mass of hydrogen is

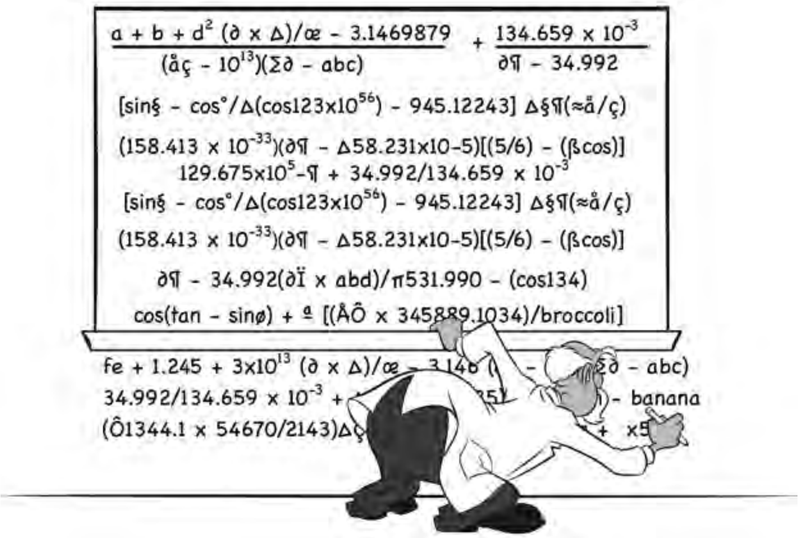
---

[10] I'm using numbers that are fairly large by chemical standards for the amounts of each substance we have, just so the math is relatively easy. In practice, these are huge amounts of zinc and hydrochloric acid, and not amounts your average nonindustrial chemist is likely to use.

1.008 and the atomic mass of chlorine is 35.45. Adding these together gives a molecular mass of 36.46 for HCl. So one mole of HCl has a mass of 36.46 grams. Because we have two moles of zinc, we would need four moles of HCl (a 2 to 1 ratio) to use up all of the zinc: 4 × 36.46 = 145.8, meaning four moles of HCl has a mass of 145.8 grams. I told you at the beginning that we had 148 grams of HCl, so we have more than enough HCl to react with all of the zinc.

6. Let's create another problem from the previous one. When you obtain hydrochloric acid, it's usually dissolved in water. Suppose the writing on the side of the container says 0.5 M. This means the *molarity* of the solution is 0.5, which means you have 0.5 moles of HCl per liter of solution. How many liters of this solution do you need to complete the above reaction? Well, we have two moles of zinc, so we need four moles of HCl; there's half a mole in each liter of solution, so we would need eight liters (!) of 0.5 M hydrochloric acid solution to complete the reaction. I told you in footnote #10 that the amounts stated in the problem were rather large!

7. In this chapter, we've dealt with ideal gases. The molecules in an ideal gas take up no space and don't have any energy due to vibration or rotation. Of course, real atoms and molecules *do* take up space and most have rotational and vibrational energy. The kinds of energy that we ignore in saying that a gas is an ideal gas affect the temperature of the gas, and the fact that real atoms and molecules take up space affects our calculations of the volume available to the gas. The corrections for a real gas are important if you're doing scientific experiments, and they can be mathematically complex. For purposes of understanding gases in general, the ideal gas is a pretty good approximation.



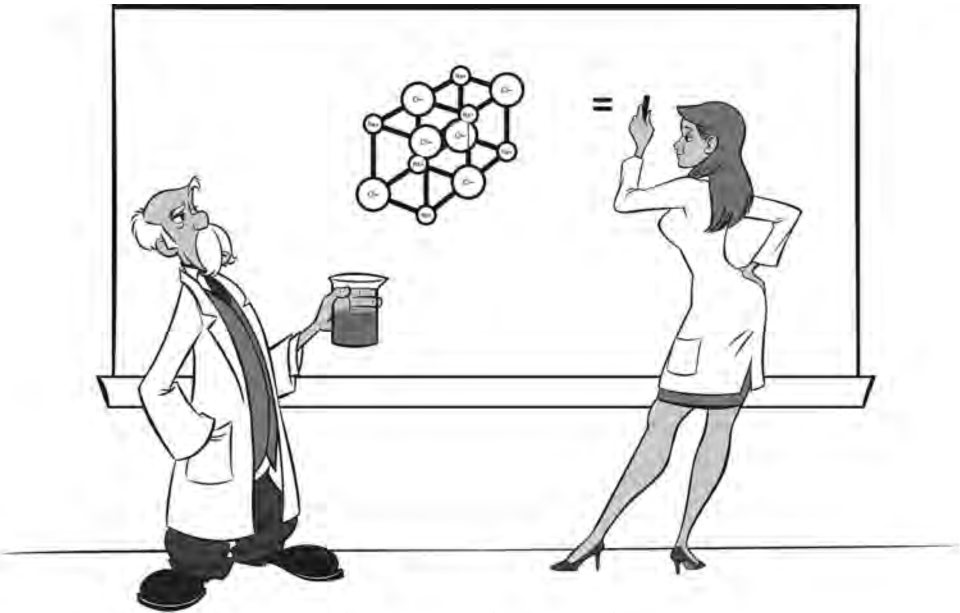
3 Chapter

The Name's Bond ... Pi Bond

You'll understand the title of this chapter once you've finished it. By the way, *pi* is pronounced *pie* and represents the Greek letter π. In the meantime, I want you to recall that we picture electrons in atoms as residing in **orbitals**—those fuzzy things that can be spherical, dumbbell-shaped, or even ring-shaped. I'm first going to address orbitals and why they're fuzzy. Then I'm going to deal with how the shapes of orbitals determine the kinds of bonds that form between atoms. Finally, we'll see how the location of orbitals and the electrons in them determine the shapes of molecules.

SCILINKS®
THE WORLD'S A CLICK AWAY

Topic: Atomic Orbitals
Go to: *www.scilinks.org*
Code: MCB009



"The name's Bond ... Pi Bond."

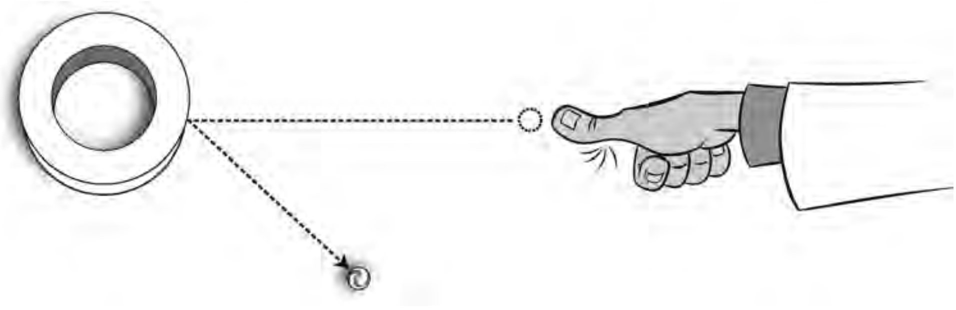
In discussing and drawing pictures of orbitals, there's potential for confusion in that you might think that if you were small enough to look at an atom directly, you would actually see these fuzzy shapes. Not so. Orbitals represent possible locations of electrons, not a physical thing. This idea is similar to drawing the orbits of the planets around the Sun. We draw those lines to represent the paths of planets, but those lines don't really exist in space. In the same way, atomic orbitals don't exist as physical things but rather are mathematical representations of where we are likely to find electrons.

Things to do before you read the science stuff

Grab something that is unlikely to move when you hit it with a marble, such as a roll of masking tape or a paperweight. Any heavy object will do, but I'll assume from here on out that you're using a roll of masking tape. Also, grab a couple of marbles and find a smooth surface.

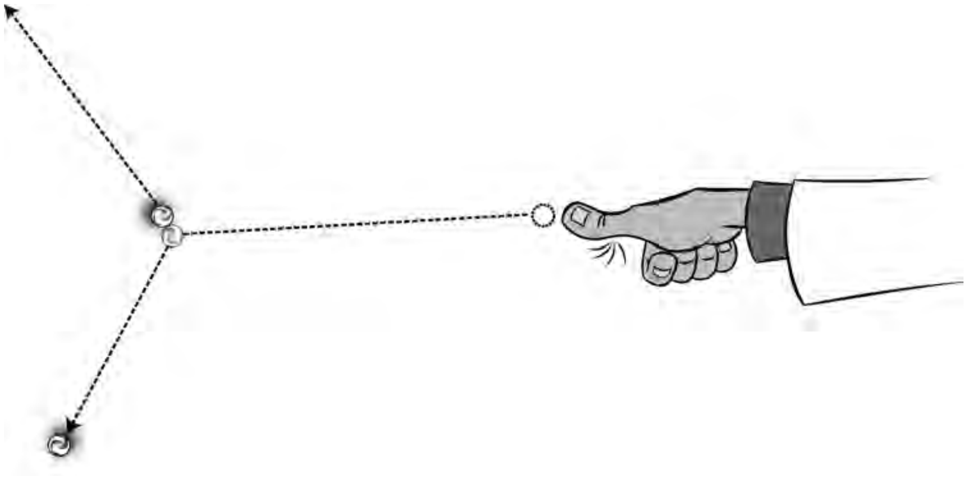
Place the masking tape on the surface, and take a trip to imaginationland, in which this roll of masking tape is invisible. Your task is to figure out exactly where this roll of masking tape is by rolling marbles at it and seeing what happens to them after they hit the masking tape. See Figure 3.1. Question: If you rolled a marble at this roll of masking tape more than a hundred times, from all different angles, could you be reasonably certain where the tape is and how big it is? (Remember, it's invisible.)

Figure 3.1



Replace the masking tape with another marble. Again, pretend this second marble is invisible. And again, roll your other marble at this "invisible" marble and see what happens to the marble you rolled. See Figure 3.2. Question: If you roll marbles at this "invisible" marble a hundred times, could you figure out how big this marble is and where exactly it is?

Figure 3.2



The science stuff

Unless you throw marbles really hard or have a wimpy roll of masking tape, I'm betting the roll of masking tape (or whatever object you used) didn't move when you rolled marbles at it. And I'm betting you reached the conclusion that if the tape was invisible and you rolled enough marbles, you would be able to determine the size and location of the tape fairly accurately.

Firing a marble at a marble, of course, is a different story. The "invisible" marble moves away when you hit it with another marble. So, while you might have known where the invisible marble *was* when you hit it, you do not know where it *is* (after the collision) or how fast it's moving. Therefore, you can't accurately determine a marble's location by hitting it with another marble. This leads us to a conclusion:

> When you try to determine an object's location by hitting it with a similar-size object, you disturb the original object's location and lose information about where it is.

What's this got to do with electrons? Be patient, and I'll get there. In the meantime, imagine you're in a room with no light and you want to know where the walls are without moving around and bumping into them. One way to determine where the walls are is to throw tennis balls at the walls. By estimating the speed at which the balls move and listening for when they hit the walls, you can eventually get a good idea about where the walls are, even though you can't see. You figure out where the walls are by hitting the walls with something. Now let's turn the lights on. You can see where the walls are. To do this, you are still hitting the walls with something. One of the scientific models for light is that it's composed of tiny particles called **photons**. According to this model, the way we see objects is by bouncing photons off of them.[1]

Because these photons are insignificantly small compared with the walls in your room or other ordinary objects, all this bouncing by millions of photons doesn't affect the shape or position of your average wall or object. This is like hitting the roll of masking tape with a marble—the marble is small enough that it doesn't significantly disturb the position of the masking tape. We can extend this notion not just to where things are but also to what they're doing; for example, how fast they're moving. The way we know where things are and what they're doing is by bouncing other things off of them. For everyday objects, we're talking about bouncing tiny little photons off them. This process is not unlike bats finding insects by bouncing sound waves (which can be thought of as tiny particles called **phonons**!) off the insects and listening for the reflected sounds. Those phonons are still quite small compared with the insects, so a bat hitting the insects with phonons doesn't disturb them very much. If the insects were of atomic size, however, a bat's sonar would be useless. The phonons hitting the insects[2] would change the position and motion of the insects so much that the bat couldn't locate the insects. Of course, if you're a bat hunting for atom-size insects, you're going to starve to death anyway.

Okay, let's move on to figuring out where electrons are and what they're doing. How are we going to do that? Well, we're going to fire things at them. Let's suppose we reflect a bunch of photons off the electrons.[3] The photons are large enough compared to electrons that when the photons run into the electrons, they disturb both the electrons' positions and their motion. So, you can't get an accurate picture of where electrons are and what they're doing because the very act of observing the electrons disturbs them. Hitting electrons with photons to find out where the electrons are is like hitting a wall with a wrecking ball to find out where the wall is. See Figure 3.3, and please realize that this is supposed to be a humorous depiction of what's going on rather than an accurate model of photons interacting with objects.

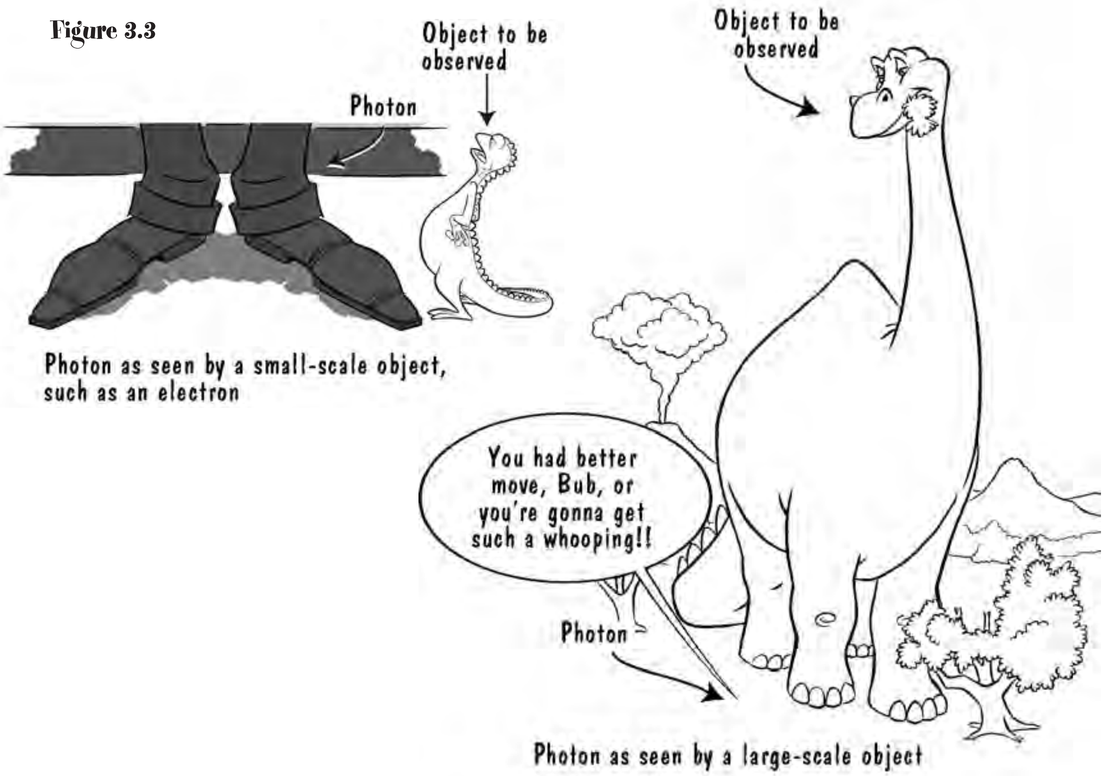
---

[1] The exception to this is when the objects around you emit their own light.

[2] Remember we're talking about sound rather than light here. So the bats are hitting the insects with sound particles (phonons) rather than light particles (photons). No, you don't need a full understanding of phonons to get this idea. You can just trust me that it's possible to think of sound as being composed of tiny particles, just as it's possible to think of light as being composed of tiny particles.

[3] We can also investigate electrons by firing particles other than photons at them, but that doesn't change the situation.

Figure 3.3



This is a general principle that applies when investigating all tiny systems such as atoms and things smaller than atoms:

> When dealing with small-scale systems, the very act of observing the systems disturbs them so you can't know precisely where the systems are or what they're doing.

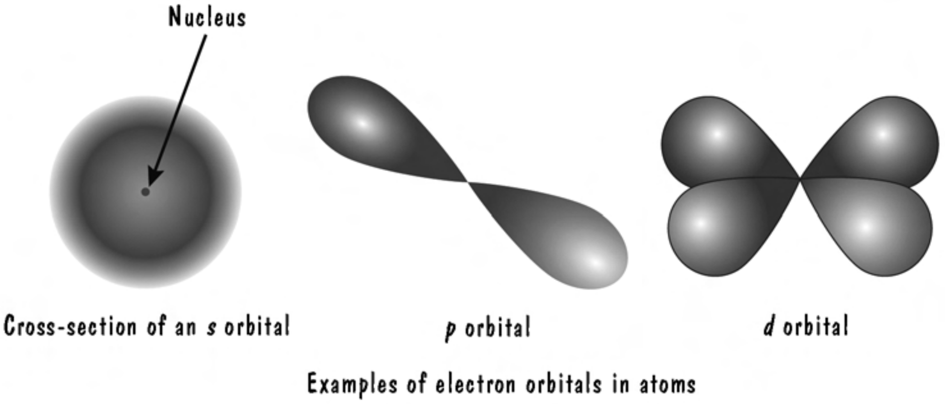
In other words, there is an *uncertainty* in what we know about the position and motion of very small objects such as electrons. In fact, you've probably heard of this principle, known as **Heisenberg's uncertainty principle**, named for the physicist Werner Heisenberg. The uncertainty principle is an integral part of the mathematics of quantum mechanics, and there is more to it than the analogy of hitting masking tape and marbles with marbles, so I don't want to give the impression that the principle is simpler than it actually is. On the other hand, our analogy is appropriate enough that you should now have some idea of why we speak of the location of electrons in terms of probability distributions[4]

---

[4] For those of you who didn't read the first chapter in this book or the first chemistry book in this series, or if you just plain forgot, fuzzy orbitals represent probability distributions, which are representations of the probability of finding an electron in any given portion of space. The darker the shading in a probability distribution, the more likely you'll find an electron in that place.

(those fuzzy orbitals) rather than specific paths or motions. We can't see the little guys, and we're hitting them with objects that are comparable in energy and size to the electrons themselves. Just as a reminder, Figure 3.4 shows a few of these electron orbitals that are probability distributions.

Figure 3.4



It's important to note that the uncertainty in measurement I'm talking about is not due to scientists just not having sensitive enough equipment. No matter how sensitive the equipment, it is a fundamental fact of nature that we cannot get exact information about what small-scale systems are doing and where they are. In other words, nature will always keep us somewhat in the dark about these measurements. This limitation has bothered many scientists for many years. Even Albert Einstein didn't accept it at first. In referring to the probabilistic nature of the atomic world, in a letter to fellow physicist Max Born, Einstein stated, "I, at any rate, am convinced that He [meaning God] does not throw dice."[5]

I've been using marbles and atom-size insects as an analogy for electrons, but I don't want to leave you with the misconception that electrons can only be thought of as solid objects. In the introduction to this book and in the first chemistry book, I discussed how we can think of electrons (and all particles, for that matter) as collections of waves. It is this wave nature of electrons that is the basis for quantum mechanics, which is the math we use to come up with the uncertainty principle. So, while it is often convenient to consider electrons to be tiny, solid objects, you should always be aware of the model of electrons as waves.

---

[5] Letter to Max Born (4 December 1926); *The Born-Einstein Letters* (translated by Irene Born). Walker and Company, New York, 1971.

I'm going to get just a bit mathematical now, so this is my official warning to those of you with math phobia. I'm not going to use this math in the rest of the book (with the exception of one application at the end of this chapter), so if you really want to skip this text box, that's okay. Anyway, we can express Heisenberg's uncertainty principle mathematically. Two inequalities cover it, and they are

$$p\Delta x \geq \hbar \text{ and } \Delta E\Delta t \geq \hbar$$

where Δp is the uncertainty[6] in measurement of momentum (the momentum of an object is its mass multiplied by its velocity), Δx is the uncertainty in measurement of position, ΔE is the uncertainty in measurement of energy, and Δt is the uncertainty in measurement of time. h is a number known as Planck's constant.

When you add the slash across the *h*, you read it as *h-bar,* and its value[7] is h divided by 2π. So, what do these inequalities mean? We'll focus on $\Delta p\Delta x \geq \hbar$. This inequality says that the more we know about a particle's position (meaning we get Δx to be very small), the less we know about its momentum (Δp must be correspondingly large to satisfy the inequality). Conversely, the more we know about the momentum of a particle (Δp is very small), the less we know about its position (Δx is very large). To restate that, the more accurately we know a particle's momentum, the less we know about the particle's position, and vice versa. To state it in even one more way, if we know almost exactly what a particle is doing, we know almost nothing about where it is; if we know almost exactly where a particle is, we know almost nothing about what it's doing.[8] Enough math reasoning. I'll address the relationship $\Delta E\Delta t \geq \hbar$ in the Applications section.

What this book is about. It might be important at this point to remind you of the purpose of this book. Reviewers of the first drafts of this book wondered why I cover things such as the uncertainty principle when the vast majority of my intended audience will never teach this material. My view is that you become a better teacher when you know more than you have to teach, so when that bright student asks a difficult question, you have some idea of how to proceed. Plus, the more complete the picture you have of what's going on at an atomic level, the better you understand the more basic stuff. Also, just because certain concepts have traditionally been considered more advanced does not mean that the average person cannot understand them. Between the discussion of probability distributions in my first chemistry book and the discussion here, I think you can get a good grasp of what electrons are doing in atoms.

---

[6] The symbol Δ means *change in* or *uncertainty in*. It is not a separate variable or constant.

[7] You will see many different versions of the uncertainty principle, each having something slightly different on the right side of the inequality. Each version has either h or $\hbar$ divided by 2 or multiplied by 2 or some such thing. The reason for the different versions is that scientists use a variety of ways to define the uncertainty in a quantity. The basic principle remains the same, regardless of the value of the constant on the right side of the inequality.

[8] In Chapter 2, I mentioned that it was impossible for any substance to reach absolute zero. The reason for this is that at absolute zero, particles are motionless, meaning we would know exactly how much momentum they have (zero) while simultaneously knowing exactly where they are. This violates the uncertainty principle, so nothing can be at absolute zero.

More things to do before you read more science stuff

Time for a couple of real activities, which will get us into the subject of atoms and their electrons in orbitals getting together. Get a sheet of wax paper and put a few water drops on the paper. Then pick up an edge of the paper so as to make separate water drops collide. What happens when they collide? Does the new shape look like a regular water drop, or is it somewhat different?

Now get yourself three small bottles of bubble solution with little plastic bubble wands. Also get yourself a friend who likes bubbles. You can do this first part by yourself if you're semicoordinated. Get a plastic wand in each hand and then blow a bunch of bubbles. Catch a bubble on each wand, as shown in Figure 3.5.

Note the shape of each bubble. Now slam the bubbles together. Do this hard enough that the bubbles join together, but not so hard that you break one or both of the bubbles. After a bit of practice, you should end up with something that looks like Figure 3.6.

Here's where the friend comes in. Have your friend blow and catch a third bubble on a third wand. Then slam this third bubble into the two joined bubbles, so you get something like Figure 3.7.

Now answer a question. Why did the bubbles change shape when you slammed them together? Think energy.

Figure 3.5

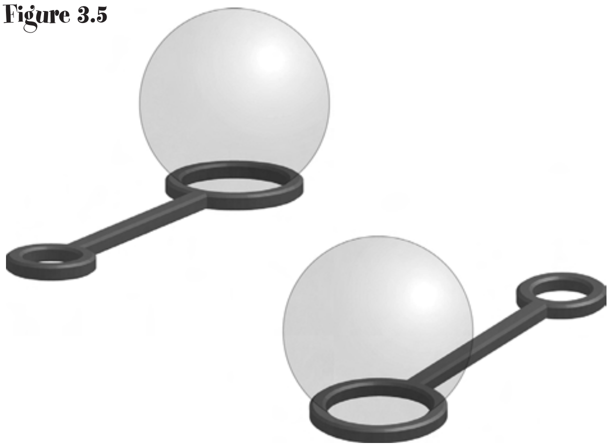


Figure 3.6

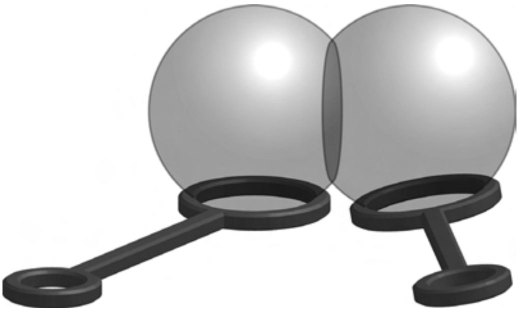
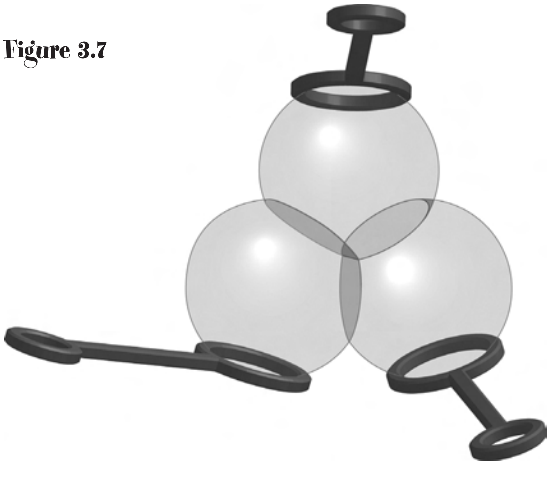


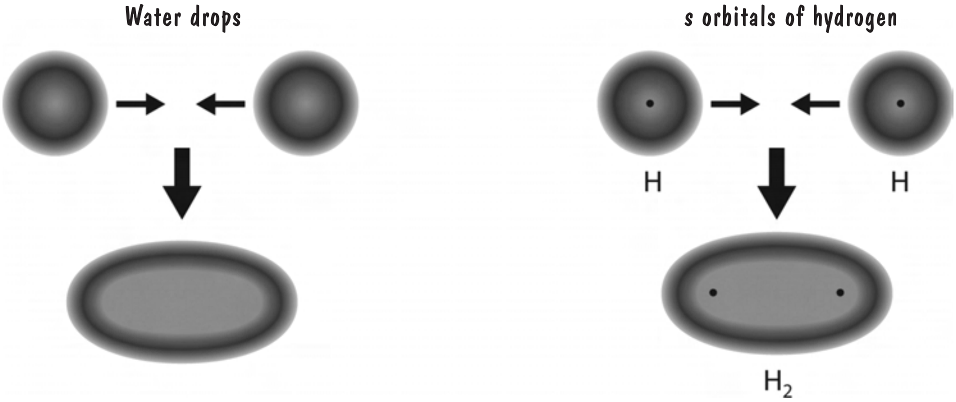
Figure 3.7



More science stuff

First let's talk about your water drops. When you bring two drops together, they join into a larger shape. This larger shape is elongated, at least at first, and is a reasonable picture of what happens when two *s* orbitals share electrons. This happens when two hydrogen atoms get together to form a hydrogen gas molecule. See Figure 3.8.

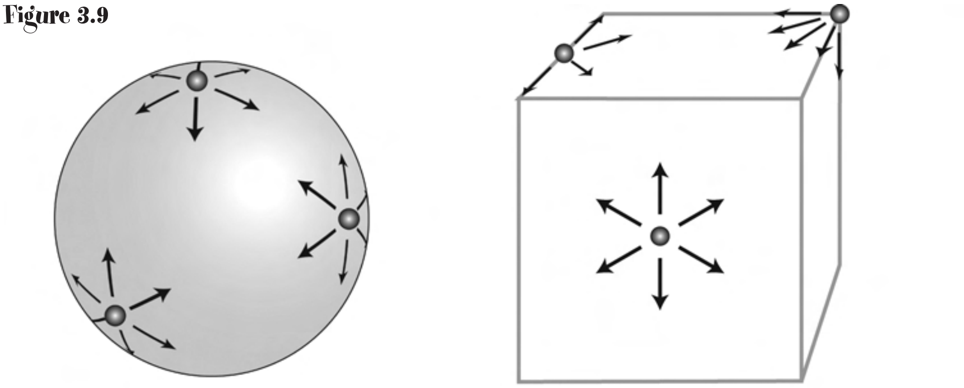
Figure 3.8



The type of covalent bond illustrated in Figure 3.8 is called a **sigma bond**. Chemists often just use the Greek letter sigma (σ) to indicate this type of bond, as in σ-bond. In sigma bonds, the orbitals connect along the axis between the two atoms. It's important to realize what's going on here. Just as the two water drops formed a new drop, the two *s* orbitals, one from each hydrogen atom, formed a brand-new orbital. This orbital is shared by the two hydrogen nuclei (which are simply protons) and contains the two shared electrons. This new orbital is known as a **hybrid orbital**.

Now, on to the bubbles and why they do what they do. This might seem like a diversion, but it will connect with what I just explained about hydrogen atoms getting together. For starters, let's figure out why bubbles are spherical in the first place. Why not a square or some other shape? The answer is that a sphere is the configuration that has the lowest energy for the bubble. Each molecule in the bubble exerts an attractive force on the molecules around it (you have undoubtedly heard of this as *surface tension*). All of these forces are equal in strength, and therefore can only be accommodated with a spherical shape. For example, if the bubble were a cube, then the forces between molecules along the sides of the square would be different from the forces between molecules at the corners. This would also be a higher energy situation overall for the bubble. See Figure 3.9 (p. 52).

Figure 3.9



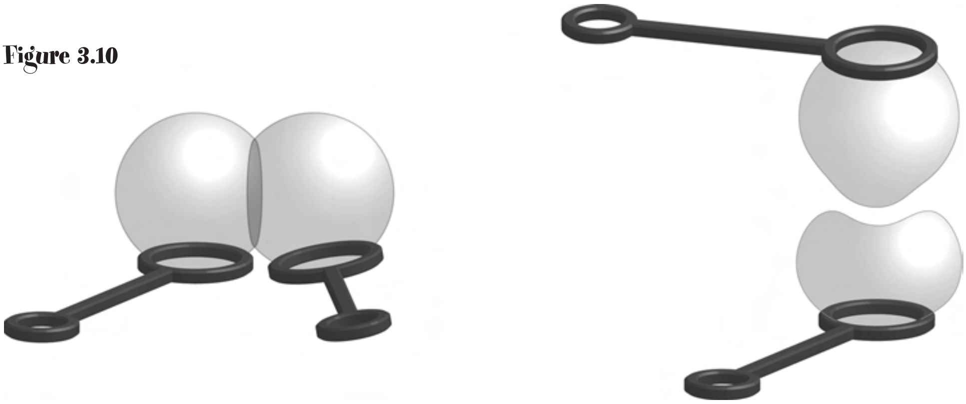
In a sphere, each molecule exerts an equal force on each other molecule. This is the lowest energy possible.

In a cube, the focus on molecules at the corners and on the edges are different from the forces on the faces. This is a high energy shape.

Now let's look at slamming the bubbles together. What you have are two things (the bubbles) that are somewhat flexible vying for the same space. The bubbles are also made of the same kind of molecules, so they can potentially interact. In other words, the molecules in one soap bubble won't differentiate between molecules that are part of their own bubble and those that are part of a different bubble. The result is the bubbles combine into a new shape that has a lower energy than, say, a couple of distorted bubbles occupying the same space. Figure 3.10 shows this distinction.

Notice that this new shape wouldn't happen if the bubbles were made of a solid substance such as rubber. If you had rubber bubbles, they would just collide and not form a new shape. The reason? It would require a great deal of extra energy to get the solids to intermingle—you would have to supply enough heat for them to melt, and then they could join together. So, it's not energetically favorable for rubber bubbles to join together to form a new shape, but it *is* energetically favorable for soap bubbles to join together to form a new shape. If we want, we could call this new shape, shown in Figures 3.6 and 3.7, a **hybrid bubble**. The hybrid bubble is not simply the overlap of the two original bubbles but an entirely new shape.

Figure 3.10



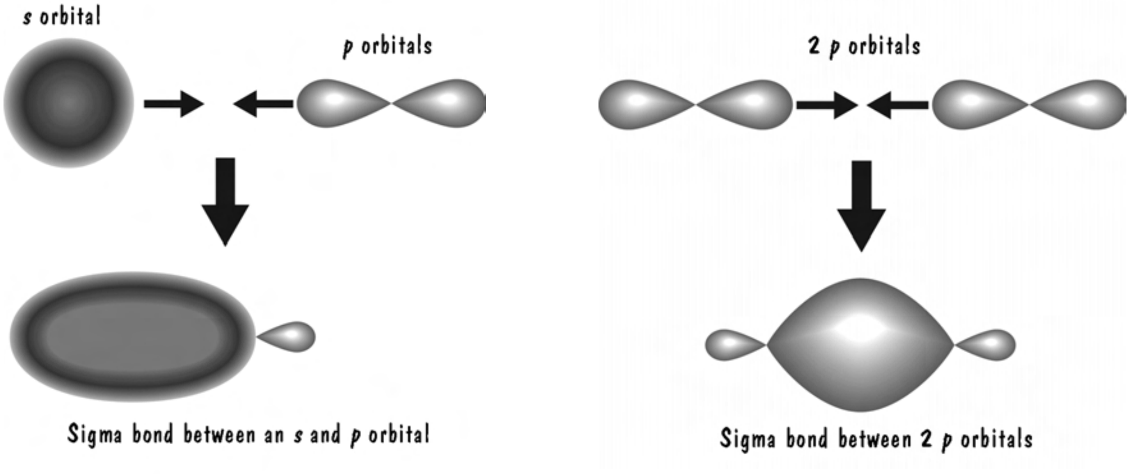
Now, what in the world does this have to do with two hydrogen atoms getting together? Well, the orbitals in these atoms are kind of like soap bubbles. In their original spherical shapes, *s* orbitals represent low energy for an electron in a hydrogen atom. In the process of overlapping and vying for the same space, these orbitals, just as soap bubbles, form into new shapes. The new shape is a lower energy than a simple overlap of the two original orbitals. There is one really important difference between bubbles and orbitals: The bubbles are real, tangible objects. Orbitals are mathematical representations.

A bit more about this new orbital (or any orbital, for that matter) being a mathematical representation. The orbital isn't real in the sense that atom-size people could see it, but rather it's a description of the mathematical probability of finding the two electrons in the space around the two hydrogen nuclei. These electrons clearly are more likely, after the joining of the two atoms, to be found in the region between the nuclei. So, the water drop and bubble analogies are just that—analogies. Orbitals are not real in the same sense that water drops and bubbles are real.

This formation of a new, hybrid orbital obviously happens when you have two atoms with partially filled *s* orbitals, but it can also happen between an *s* orbital and a *p* orbital or between two *p* orbitals. Check out Figure 3.11. Both of the bonds represented there are sigma bonds because they form on the axis between the two atoms. They also represent the sharing of electrons between the two atoms in the newly formed hybrid orbital.

Notice in Figure 3.11 that the original orbitals change shape in a couple of ways. First, much of the space between the nuclei is taken up with orbitals, and

Figure 3.11



the outside part of the original orbitals (for example, the lobes on the far right and left in the *p-p* combination) gets smaller. What that means is that after the bond is formed, the electrons shared by the atoms spend more of their time between the nuclei than in the spaces not between the nuclei. Makes sense if those electrons are shared electrons. See Figure 3.12.

Before moving on, take a moment and ask yourself why two atoms get together in a covalent bond in the first place. Don't negative charges repel one another? When two partially filled orbitals get near each other, shouldn't they push apart, as in Figure 3.13?

Well, electrons have a property I haven't yet discussed. This property is known as the **electron spin**. The spin of an electron can be up or down. Electron spin is just one more quantum number that one assigns to electrons, which I've discussed earlier. When two electrons with the same spin (both up or both down) come together, they repel each other. When two electrons with opposite spins (one up and one down) come together, there is an attractive force. This attractive force is stronger than the electric repulsion between the electrons, so opposite-spin electrons get together and share the same orbital (in the case of electrons within a single atom) or share a combined hybrid orbital (in the case of electrons from separate atoms sharing a combined orbital). See Figure 3.14.

Figure 3.12

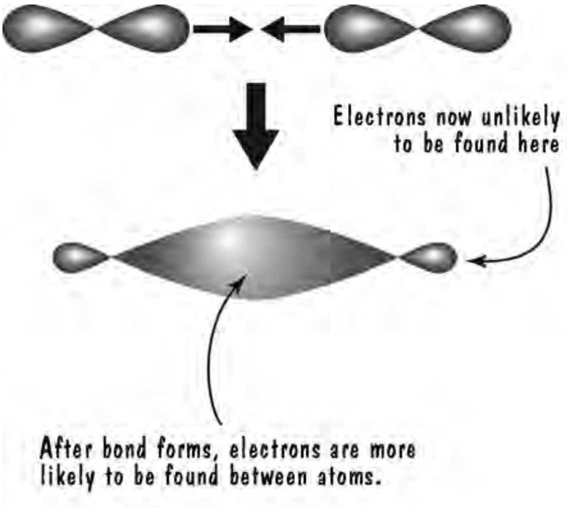
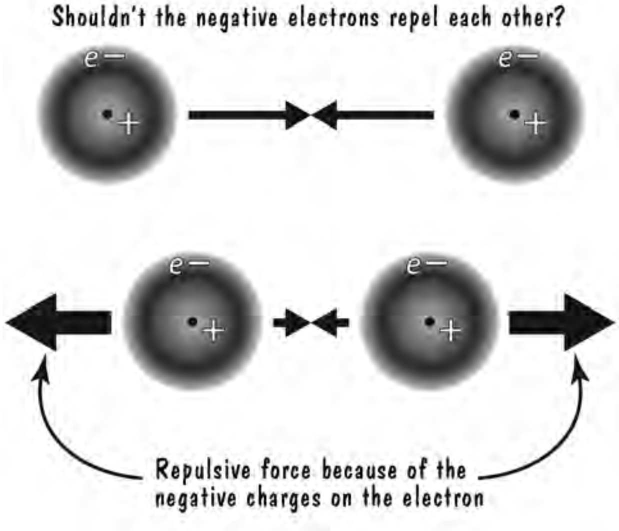
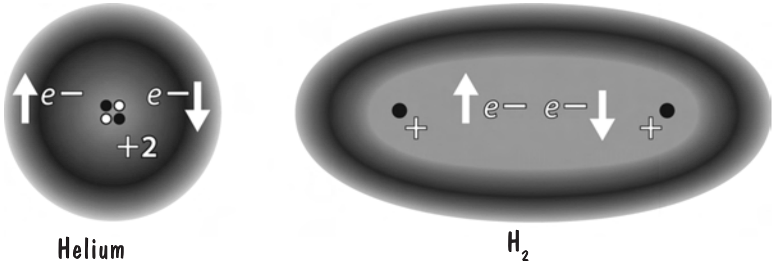


Figure 3.13



Topic: Electrons in an Atom
Go to: *www.scilinks.org*
Code: MCB010

Figure 3.14

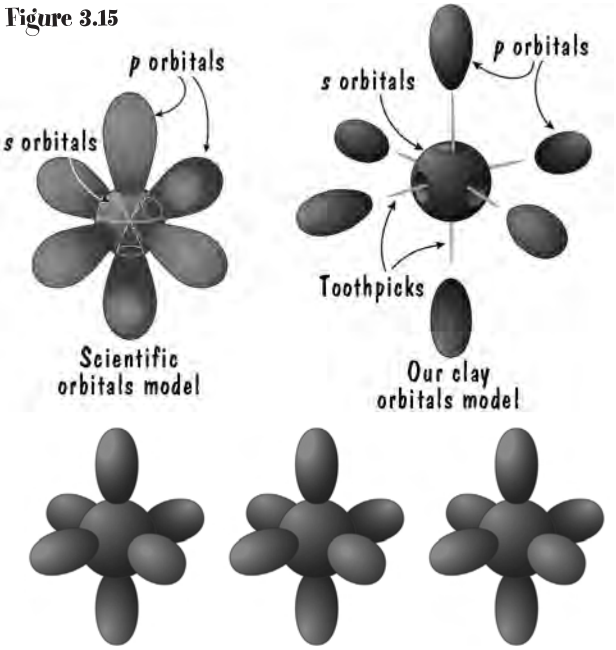


Electrons with opposite spins (indicated by arrows) have an attraction that keeps them in the same orbital despite electric repulsion.

So, electrons tend to repel one another because they are both negatively charged. Electrons with opposite spins have an attractive force that overcomes the electric repulsion. Thus, opposite-spin electrons from different atoms will form a bond, but electrons that are already paired up by spin, as happens in a filled orbital, will repel other pairs of electrons. This means that filled orbitals repel one another, which will come into play in just a bit.

Even more things to do before you read even more science stuff

Figure 3.15



Now get a bunch of toothpicks and some modeling clay. What you're going to create are a few atoms that have electrons in a filled *s* orbital and partially filled *p* orbitals. (Refer back to Chapter 1 if you need to review what I mean by *filled* and *partially filled* orbitals. Full *s* orbitals contain two electrons, and full *p* orbitals contain two electrons, leading to six total electrons in three filled *p* orbitals.) Using toothpicks and modeling clay, build three of the structures shown in Figure 3.15.

These structures represent the *s* and *p* orbitals in the outer shell of an atom.[9] This atom could be nitrogen, as that's an atom that

---

[9] The sizes of the *p* orbitals in these drawings are greatly exaggerated, but that exaggeration is necessary for you to visualize things in this activity.

has a filled outer *s* orbital and partially filled outer *p* orbitals. In other words, each dumbbell-shaped *p* orbital has one electron in it, so each *p* orbital is half full. See Figure 3.16.

Figure 3.16

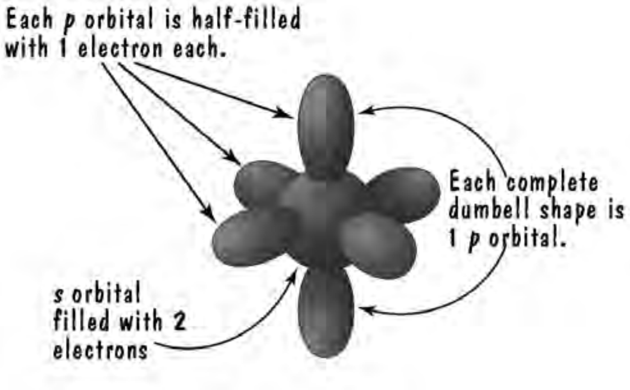


Figure 3.17

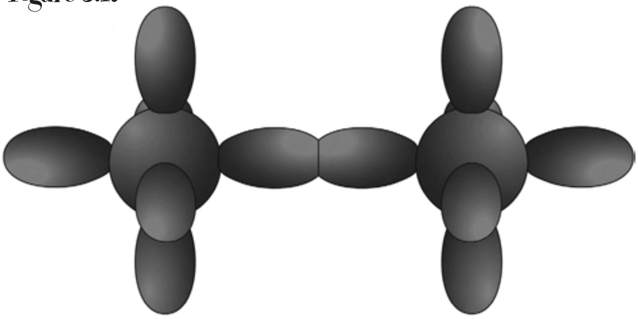
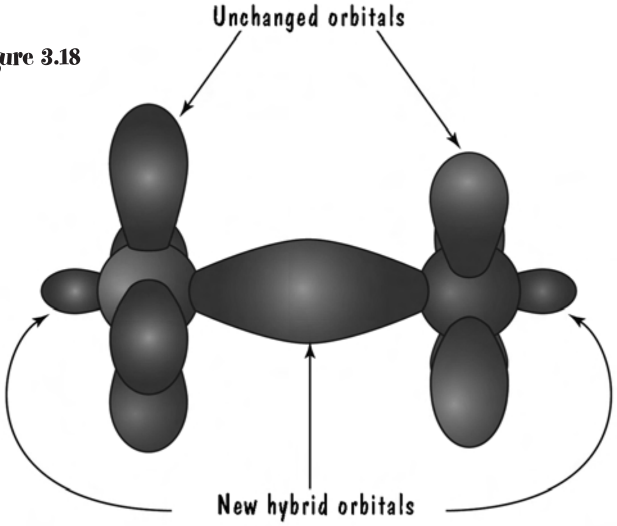


Figure 3.18



Bring two of these atom models next to each other. As you move the models closer together, ask yourself which orbitals will come in contact first, second, and which third. Assume that when orbitals from different atoms touch, then the orbitals begin sharing electrons in a covalent bond. Also, remember that our *s* orbitals are filled in this example, so there is no sharing of electrons between the *s* orbitals. To get you started, Figure 3.17 shows the two orbitals that will get together first, at least in our model. There's nothing special about the particular *p* orbitals shown connecting with each other. The point is that the "extremities" of the atom will come in contact first, as people with outstretched arms are likely to touch other people first with their hands rather than their bodies.

Once that covalent bond forms, those two connecting *p* orbitals are going to change into a new hybrid orbital. That means, to have a more accurate model of what might actually happen with orbitals, we have to alter our clay to look like Figure 3.18.

Given that the hybrid orbital you just formed stays put, try to figure out how the other *p* orbitals might get together. This will obviously form two more covalent bonds, or what we call a **triple bond** between the atoms. The next explanation section will clarify this idea further.

Now go back to your three original clay models. If you altered your models to resemble Figure 3.18, change them back. Using these three models, and assuming they connect together in only sigma bonds, try to figure out what shape molecules these atoms can form—spheres, cubes, pyramids, or something else.

Even more science stuff

Now let's see how you did with your clay structures. I already showed you the first orbitals to meet, which are *p* orbitals meeting head on. This is a sigma bond. Once these orbitals are sharing electrons in a covalent bond, you can rotate one or both of your atoms so the other *p* orbitals are lined up, as shown in Figure 3.19.

Then you can see how the other *p* orbitals can form sort of a "sideways" bond, pictured in Figure 3.20.

The "sideways" bonds shown in Figure 3.20 are called **pi** (pronounced like *pie*) **bonds**. Sigma bonds are stronger than pi bonds, which leads us to the following rule:

When atoms get together, the first bond is always a sigma bond.

Once that bond is formed, then pi bonds can form. Figure 3.20 shows a triple bond between two atoms. For example, when two nitrogen atoms get together to form N_2, there is a triple bond (a total of six shared electrons) between the two atoms—one sigma bond and two pi bonds. When atoms get together and form a double bond, there's one sigma bond (the one that forms first) and one pi bond.

Figure 3.19

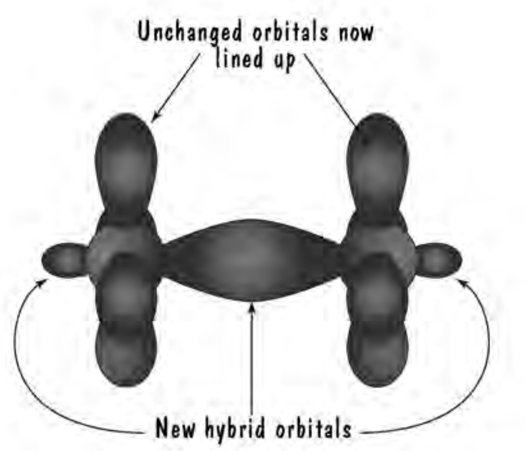
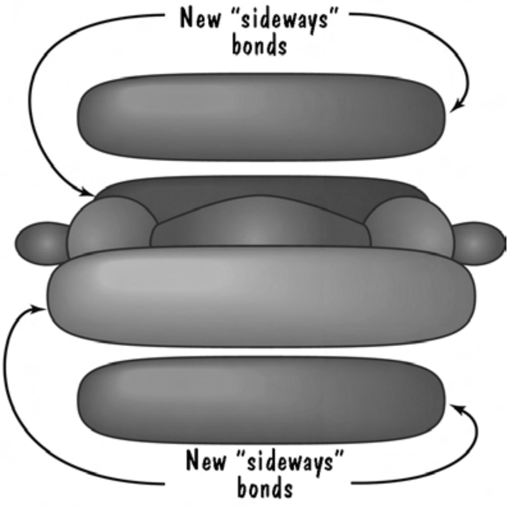
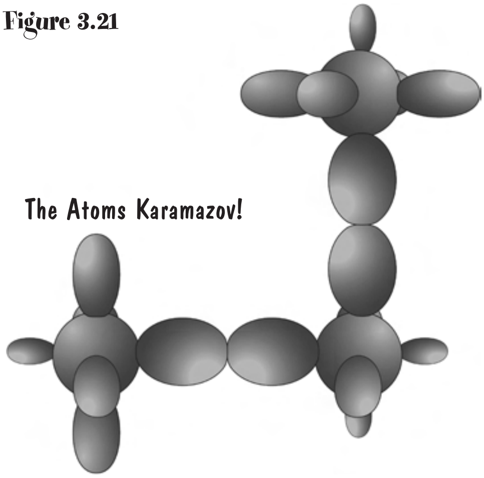


Figure 3.20



That's all well and good, but what happens when multiple atoms get together? I had you simulate this event in the second part of the previous section. If you allow only sigma bonds to form between your clay atoms, then you likely came up with a formation like the one in Figure 3.21.

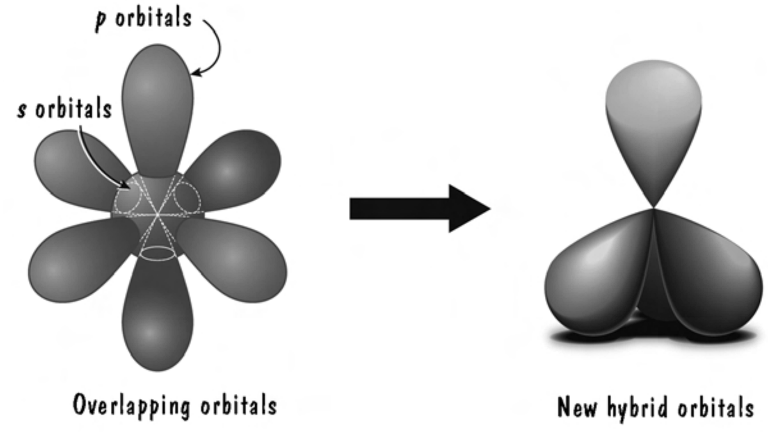
Figure 3.21



Now imagine adding more and more clay atoms to this structure. Because of the 90-degree angles in the structure, what you're going to form is a cube. If this holds true for all atoms, then that would mean that all molecules (collections of bonded atoms) will have 90-degree angles between atoms, so that the only possible shapes for molecules would be cubes or rectangular prisms. In nature, however, we find all sorts of molecular shapes. That must mean that we don't have a complete picture of bond formation.

The answer to our dilemma lies in the formation of more hybrid orbitals. This time, though, the hybrid orbitals form within a single atom rather than between two atoms. Up to this point, I've had you consider the *s* and *p* orbitals of an atom as if they simply overlap one another. But really, we have the same situation as we did with water drops, soap bubbles, and orbitals from separate atoms coming together. The *s* and *p* orbitals within an atom are competing for the same space. So, often, instead of remaining as separate overlapping orbitals, the orbitals combine into new hybrid orbitals. Figure 3.22 shows a common form of hybridization, in which the one *s* orbital combines with the three *p* orbitals to form four equivalent orbitals that are in the shape of a tetrahedron. This is known as an sp^3 hybridization. The name results from the fact that one *s* orbital combines with three *p* orbitals.

Figure 3.22



Let's see how this kind of hybridization works for the element carbon. By looking at the Periodic Table, you can see that carbon has four valence electrons—two in the *s* orbital and two in the *p* orbitals. Because the two electrons in the *s* orbital are already paired up, you would expect carbon to share only the two *p* electrons with other atoms, as shown in Figure 3.23.

Figure 3.23

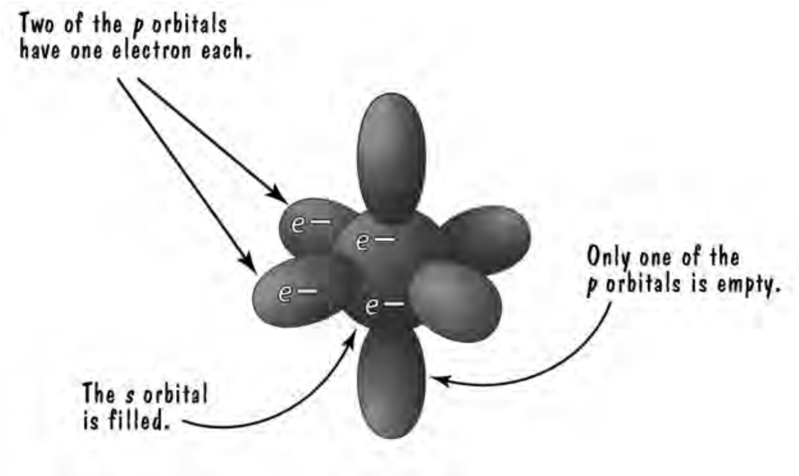
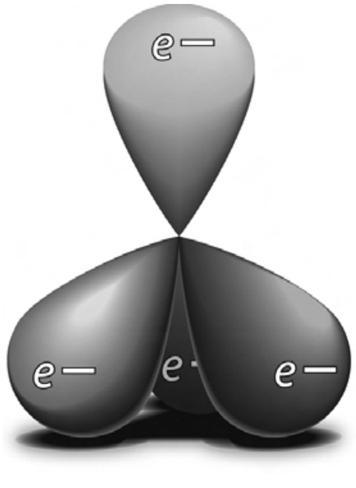


Figure 3.24

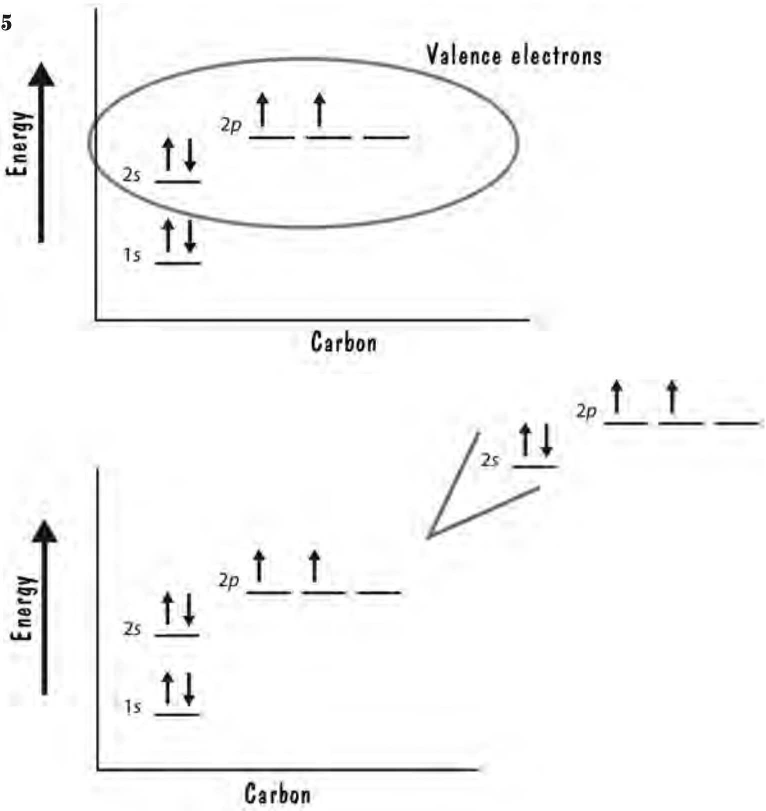


Each hybrid orbital has one electron, and those electrons are as far from one another as possible.

But the outer *s* and *p* orbitals in carbon form four equivalent hybridized orbitals (sp^3), as shown in Figure 3.24. In this arrangement, there are now four unpaired electrons that can bond with other atoms. This leads to carbon being a special atom that can form all kinds of long chain molecules. You might also notice that with this hybridization, the valence electrons in carbon are as far from one another as possible. With the electrons as far away as possible from one another, this is clearly a lower energy situation than the one shown in Figure 3.23.

Another way to look at what's happening with hybrid orbitals is to study energy level diagrams. We'll stick with what happens with carbon. In the outer shell, there are two paired electrons in the *s* orbital and two unpaired electrons in the *p* orbitals. See Figure 3.25 (p. 60).

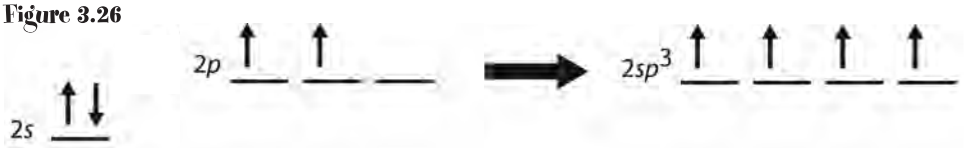
Figure 3.25



> **Important note.** I'm using a new notation (different from the first chemistry book and the first chapter in this book) to show electrons in different energy levels. Instead of solid circles, I'm now using up and down arrows. The arrows represent the electron spin, which is either up or down. This is the conventional way of doing things, as it shows the pairing of electrons with up and down spins in a given orbital.

These orbitals tend to combine into four hybridized orbitals—the sp^3 orbitals. These new orbitals all have the same energy, as shown in Figure 3.26.

Figure 3.26

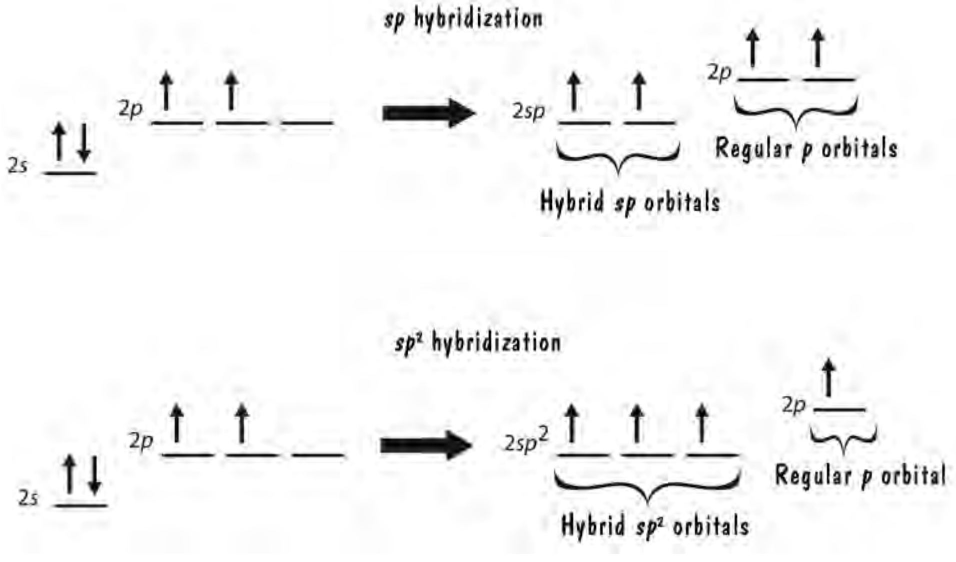


So, instead of one *s* orbital and three *p* orbitals, we have four equivalent sp^3 orbitals.

Hybridization of the orbitals has therefore caused two things to happen. First, the orientation of the outermost orbitals has changed from the shape of a cube, which would lead to 90-degree bonds with other atoms, to the shape of a tetrahedron, which leads to approximately 110-degree bonds with other atoms. Second, hybridization has also changed carbon from an atom that has three orbitals willing to share electrons to an atom that has four equivalent orbitals willing to share electrons. In that process, carbon changes from an atom with limited options for molecule shapes to one that has many options for molecule shapes. In other words, carbon is a special atom primarily because of its hybrid orbitals.

The sp^3 hybrid orbitals aren't the only ones possible. Often the *s* orbital will interact with only one or two of the *p* orbitals and leave the remaining ones alone. Then you get hybrid orbitals labeled sp^2 or *sp*. Figure 3.27 shows the change in energy levels for these two hybrids.

Figure 3.27



The *d* orbitals in an atom also take part in hybridization, leading to hybrid orbitals such as dsp^3 or d^2sp^3. There is a definite relationship between the structure of molecules and the hybridization of orbitals. Table 3.1 (p. 62) lists various hybrid orbitals and the resulting structures.[10]

---

[10] You can use this table as a general guideline, but the molecular shapes that result in any molecule will not always correspond exactly to these shapes. And no, I don't expect it to be obvious why these hybridizations lead to the shapes given. The table is here for completeness more than anything else.

Table 3.1

Hybridization	Structure
sp	linear
sp^2	planar triangle
sp^3	tetrahedron
dsp^3	trigonal bipyramid
d^2sp^3	octahedron

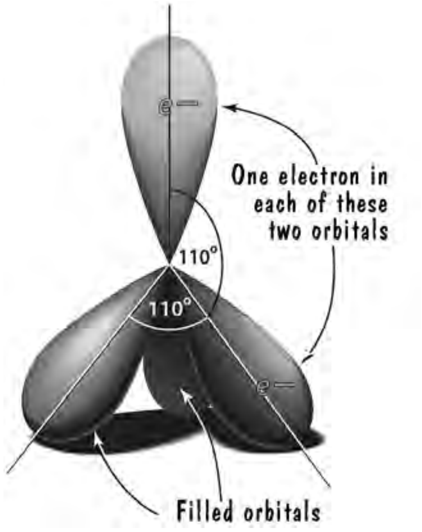
And even more things to do before you read even more science stuff

This is a short section in which I want you to just think about something. I mentioned in footnote #10 (p. 61) that there isn't always a direct correlation between hybridization and molecular shape. For example, oxygen atoms form sp^3 hybrid orbitals. You would expect, then, that oxygen atoms combining with other atoms as molecules would have an overall shape of a tetrahedron. The angles between bonds in a tetrahedron are about 110 degrees (actually 109.5 degrees). Yet, in a water molecule in which oxygen combines with two hydrogen atoms, the angle between the bonds is only 105 degrees. What do you suppose leads to this nontetrahedral shape? Think about electrons repelling one another.

And even more science stuff

To be honest, I didn't expect you to figure out that dilemma on your own, but I thought I'd give you a try. To see what's happening, let's look at the tetrahedral shape that results from sp^3 hybridization. Oxygen has six valence electrons rather than the four that carbon has. In distributing these six valence electrons among the four hybrid orbitals, you end up with two filled orbitals (two electrons each) and two half-filled orbitals (one electron each). The half-filled orbitals are the ones that bond with hydrogen atoms. See Figure 3.28.

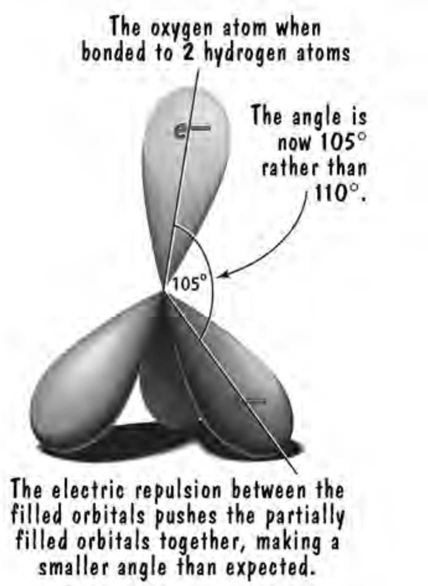
Figure 3.28



Those negatively charged filled orbitals (remember, they have electrons in them) repel each other and the other orbitals. Because the filled orbitals are larger, in general, than the bonding orbitals, and because they contain two electrons instead of one, their electric repulsion is larger. This bends the tetrahedral shape, as shown in Figure 3.29. That results in the bond angle in a water molecule being less than the expected 110 degrees.

The stronger repulsion of filled orbitals affects the shape of many molecules. That means that Table 3.1 seldom gives you the proper shape of a molecule. In addition to knowing what hybridization occurs, you have to factor in the alteration of that shape due to repulsion of filled orbitals. The water molecule is a relatively simple example, but the same principle applies in more complicated situations.

Figure 3.29



Chapter summary

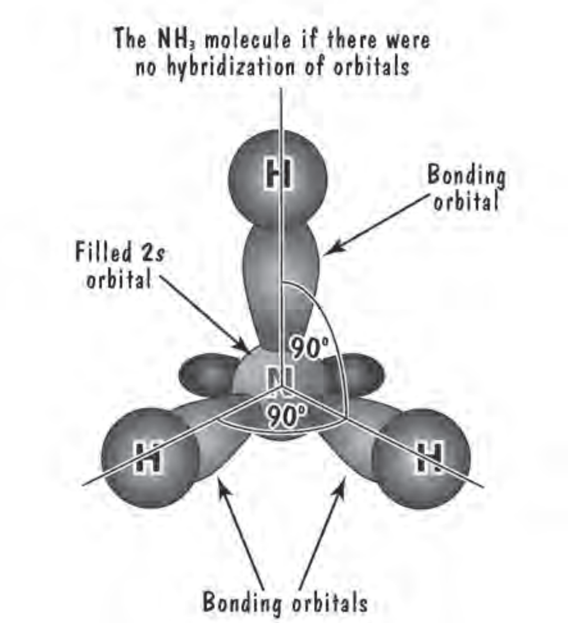
- The very act of observing what's going on at the scale of atoms and smaller things disturbs what you are observing. This leads to an inherent uncertainty in our understanding of where electrons are and what they're doing. The name given to this concept is Heisenberg's uncertainty principle.
- The shape and orientation of electron orbitals determine how atoms connect in covalent bonds and the overall structure of molecules.
- When two atoms form a covalent bond, the orbitals can connect via a head-on bond, known as a sigma bond, or a sideways bond, known as a pi bond. The first bond to form between atoms is always a sigma bond. After a sigma bond forms, then pi bonds can form. Sigma bonds are stronger than pi bonds.
- Different orbitals will combine into a hybrid orbital when that process is energetically favorable. The hybrid orbital has a different orientation and shape from the original orbitals. The simplest form of hybrid orbital is the combination of two *s* orbitals from different atoms. The existence of hybrid orbitals leads to bond angles other than 90 degrees and allows for a wide variety of molecular structures.
- Hybrid orbitals within a single atom are labeled according to which orbitals are involved. For example, an sp^2 orbital is the combination of one *s* orbital with two *p* orbitals to make three sp^2 orbitals that all have the same energy.

- Filled orbitals (containing two electrons each) tend to distort the shape of a molecule because of the stronger electric repulsion caused by these filled orbitals.

Applications

1. **Note:** You can ignore this application if you didn't read the text box dealing with the mathematics of the uncertainty principle (p. 49). I told you I would address $\Delta E \Delta t \geq \hbar$, one of the inequalities associated with the uncertainty principle, in this section. Guess I'd better do that. Some atoms are radioactive, meaning that they decay into a different form of the atom or a different atom altogether. In the process, the atoms release various subatomic particles. For example, carbon-14 spontaneously decays into carbon-12 with the release of neutrons. The relationship $\Delta E \Delta t \geq \hbar$ tells us that the more we know about the energy of the escaping neutrons, the less we know about when the decay happened. Conversely, the more we know about when the decay happens, the less we know about the energy involved. Just as with electron orbitals, this makes us speak in terms of probabilities. We can state the probability that a given atom of carbon-14 will decay in a given period of time, but we can't know exactly when it will happen. For an interesting paradox that uses this form of the uncertainty principle, search the internet for "Schrödinger's cat."
2. One thing that makes carbon such a versatile atom is its ability to form three different kinds of hybrid orbitals—sp^3, sp^2, and sp. When carbon atoms are connected in a tetrahedron, as happens with sp^3 hybridization, you get diamond. Using sp^2 hybridization, carbon forms into graphite, which has a different shape from diamond and certainly has different properties. When bonded with two oxygen atoms, carbon uses sp hybridization to form a linear carbon dioxide (CO_2) molecule.
3. Here's another example of figuring out the shape of a molecule from knowing what kind of hybrid orbitals it has. We'll look at the molecule NH_3. Nitrogen,

Figure 3.30

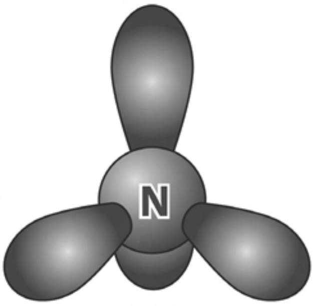


as you can tell from looking at the Periodic Table, has five valence electrons—two electrons in the 2*s* orbital and three electrons in the three 2*p* orbitals. Given this, you might think that the two 2*s* electrons would be perfectly happy to keep to themselves and let the 2*p* electrons bond with other atoms. If that were the case, then the NH_3 molecule would look like Figure 3.30, with a 90-degree angle between the bonds.

But the NH_3 molecule doesn't have 90-degree angles between the bonds, and here's what is happening. Just as with carbon, the valence *s* and *p* orbitals form four equal hybrid orbitals, as shown in Figure 3.31.

Figure 3.31

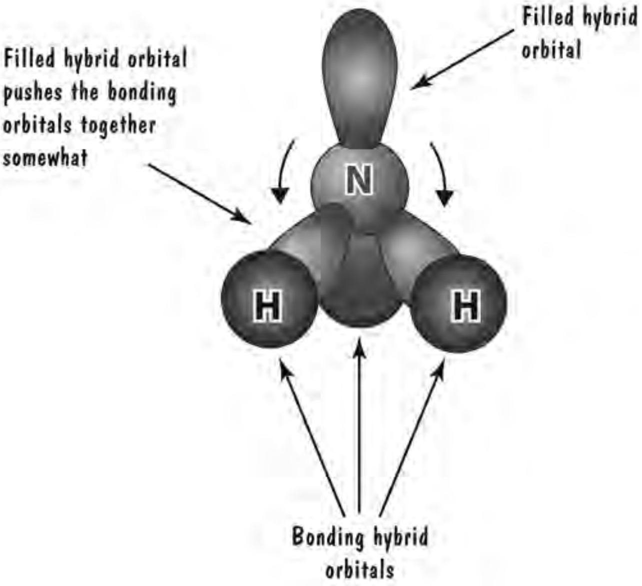
sp^3 hybrid orbitals



Unlike carbon, which has four valence electrons to be distributed among the four equal hybrid orbitals, nitrogen has five electrons to fill those orbitals. That means one of the orbitals has a complete pair of electrons. The extra repulsion that results from that complete pair of electrons pushes the other bonds slightly away. If there were no such extra repulsion, the angle between the hybrid orbitals would be 109.5 degrees (refer back to the discussion of water molecules in this chapter). With the repulsion, the angle between orbitals is more like 107.3 degrees. See Figure 3.32. Now, you might be wondering what the big deal is. Why worry about 2 degrees of angle? Well, those angles are extremely important for figuring out how and why molecules interact. Often there has to be a certain "fit" between molecules, and the proper angle can be crucial for the right fit.

Figure 3.32

Actual shape of NH_3 molecule due to hybridization and due to repulsion of filled orbital



4. In this chapter, I mentioned that atoms such as oxygen

and nitrogen can form double and triple bonds. The first bond that forms is a sigma bond. Once this is formed, other orbitals can overlap "sideways" in a pi bond. Once you get to the fourth row, however, the atoms there tend not to form double or triple bonds with other atoms. Why? The answer is that in this row, beginning with the element zinc, atoms have electrons in filled *d* orbitals. These electrons in the filled (two electrons each) *d* orbitals tend to get in the way of *p* electrons that would otherwise form pi bonds. No pi bonds, no double or triple bonds. Of course, unfilled d orbitals can easily form double and triple bonds, and if the filled *d* orbitals become involved with *s* and *p* orbitals in hybrid orbitals, those hybrid orbitals can then form pi bonds and thus have double and triple bonds. And no, this last application isn't high on the list of things you absolutely have to know to understand bonding. Just something extra for you.

4
Chapter

Special Reactions

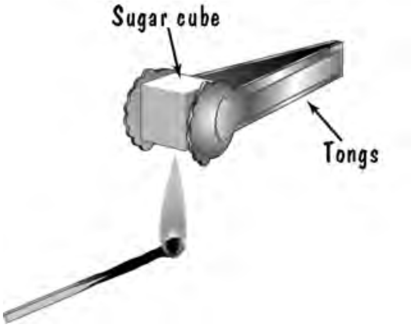
I covered the basics of chemical reactions in the first chemistry book, including how to write and balance chemical equations that represent those reactions. There is also a quick review of chemical reactions in Chapter 1 of this book. We're going to delve a bit deeper into certain kinds of reactions—ones that occur often enough that it's useful for them to have their own special categories. As you go through this chapter, try to avoid seeing these separate categories as conceptually different from other reactions. The same basic principles govern all chemical reactions, regardless of the category.

Things to do before you read the science stuff

Here's a neat trick to amaze friends and family members. Challenge someone to make a sugar cube catch on fire. For safety's sake, hold the sugar cube with tongs and place a lit match or a candle under it. See Figure 4.1.

Although your chosen person might get the sugar cube to melt or caramelize, he or she will be out of luck when it comes to the cube going up in flames. You come to the rescue with a new sugar cube. Before you place the cube in the tongs, dust it on all sides with cigarette ashes.[1] When you place your dusted cube over a flame, it should catch fire easily. Neat.

Figure 4.1



The science stuff

What you just did is an example of a **catalyzed reaction**. A **catalyst** is something that helps a reaction occur or increases the rate of the reaction. Although catalysts affect a reaction, they themselves are not affected by the reaction as a whole. The particular kind of reaction you catalyzed when getting the sugar cube to burn is known as **combustion**, which is just a fancy name for something burning. When a substance burns, oxygen combines with carbon and hydrogen atoms to form carbon dioxide and water. The products—carbon dioxide and water—are always the same in combustion. Therefore, we have a classification known as combustion reactions. All of them are of the form

organic molecule + oxygen → carbon dioxide + water

The term **organic molecule** refers to molecules that contain primarily carbon and hydrogen.

The above reaction doesn't include the fact that you often have to add energy (via the flame) to get the reaction to go forward. For our sugar cube (composed of sucrose), we can write this as

sucrose + oxygen + energy → carbon dioxide + water

or using chemical formulas (and balancing the equation):[2]

---

[1] For the record, we here at the Stop Faking It! Institute of Higher Learning do not promote smoking. I have to admit, though, that I love the smell of a good cigar.

[2] Those letters in parentheses—(s), (g), and (aq)—refer to the state of the chemical involved in the reaction. These symbols refer to solid, gas, and aqueous, respectively.

$$C_{12}H_{22}O_{11}\ (s) + 12O_2\ (g) + \text{energy} \rightarrow 12CO_2\ (g) + 11H_2O\ (g)$$

Whenever you have to add energy to get a reaction to occur, that energy is referred to as the **activation energy**. We can represent the activation energy with a graph of the energy involved in the reaction versus time, as shown in Figure 4.2.[3]

You can think of the activation energy shown in Figure 4.2 as an "energy hump" you have to get over for the reaction to proceed. Imagine the graph is a track on which you can place a bowling ball. For the bowling ball to go from the energy level on the left side to the lower energy on the right side (you eventually get energy *out* of this process because the ball is now rolling faster on the right side), you have to push the ball (add energy) up to the top of the hump. See Figure 4.3.

Now let's apply this idea to the sugar cube. For the cube to burn, you have to get over the energy hump to the right side of the graph. A match simply can't provide enough energy to do this. You can solve the problem by

Figure 4.2

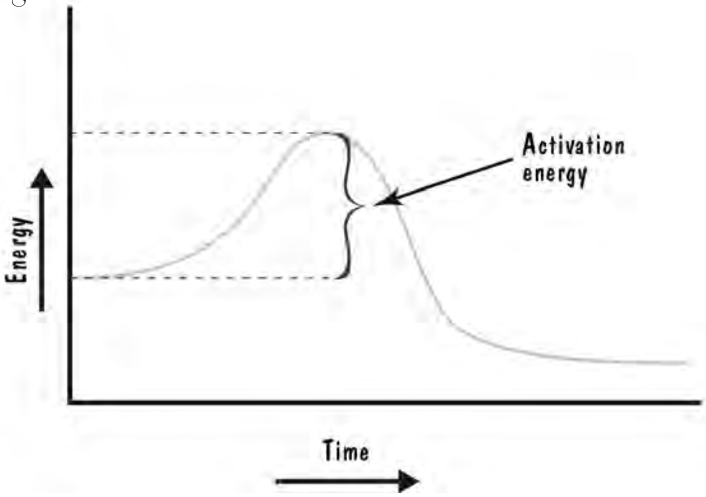
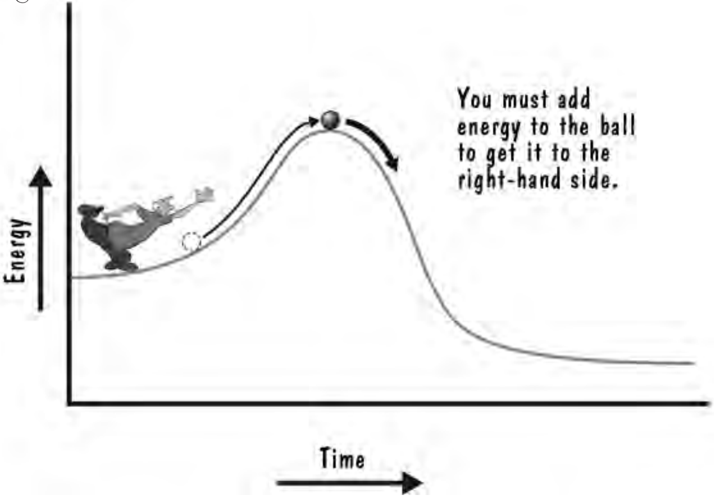


Figure 4.3

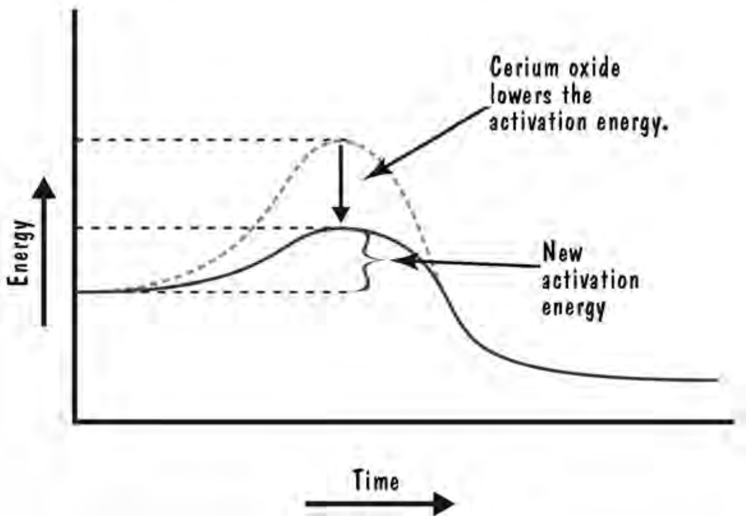


---

[3] There's potential for confusion in this graph because it looks similar to energy level diagrams for individual atoms, such as Figure 1.9 (p. 9) in Chapter 1. The vertical axis is energy in both cases, but the horizontal axis in Figure 4.2 is time, while the horizontal axis in Figure 1.9 has no meaning. In Figure 4.2, we are tracking the energy of the molecules involved in a chemical reaction as the reaction proceeds, and in Figure 1.9 we are illustrating the energy levels available to electrons in an atom.

adding more energy than a match can provide.[4] Alternatively, you can *lower* the activation energy, which is where the cigarette ash comes in. The ash contains a substance called cerium oxide. The cerium oxide acts as a go-between for the oxygen molecules in the air to combine with the sucrose molecules. With cerium oxide as the mediary, the energy required for the combustion of sucrose is lower than without the cerium oxide, low enough that the energy provided by a lit match is enough for the reaction to proceed. See Figure 4.4.

Figure 4.4



All catalyzed reactions use this same principle. By adding a catalyst to a reaction, you lower the activation energy necessary for the reaction to occur. You might think of the catalyst as a facilitator. You have two molecules that have trouble getting together on issues. In comes the catalyst as mediator, helping the two sides get together. As catalysts often say, "Can't we all just get along?"

More things to do before you read more science stuff

In this section and the next one, I'm not going to deal with reactions, but rather with chemicals that we classify as acids and bases. This little aside is necessary because following these two sections I'm going to deal with acid-base reactions. Seems silly to cover that before you know what acids and bases are! Anyway, for this section you'll need a supply of litmus paper (both red and blue) and a supply of pH paper. You can get these from a science materials supplier or possibly at a place that sells products for swimming pools or spas. You'll also need various liquids to test. Good ones that you can find around the house are vinegar, bleach, black coffee, ammonia, water, baking soda (dissolved in water), and lemon juice. Chemicals that are available from a science supply outfit or maybe already lying around your school include sodium hydroxide, calcium hydroxide, hydrochloric

---

[4] For example, you might get a sugar cube to go up in flames if you trained a blowtorch on it.

acid, and sulfuric acid. If you have these chemicals in different concentrations (different molarities)[5], that would be good.

Caution: Commercial chemicals such as hydrochloric acid and sodium hydroxide can come in concentrations that are downright dangerous. When using these chemicals and somewhat noxious household chemicals such as ammonia, always wear goggles and gloves and have a plan for washing out your eyes if they come in contact with the chemicals. Also, **take care not to mix any of the liquids you have**. Mixing ammonia and bleach, for example, will generate a dangerous gas. Not good (I inadvertently did this once and it's scary!). See Safety Note (p. ix).

Okay. First, dip strips of blue litmus paper in the various liquids you have. Use a new strip for each liquid. The paper will either stay blue or turn red. Then dip strips of red litmus paper in the various liquids. The paper will either stay red or turn blue. Keep track of what liquids cause what reactions in the litmus paper.

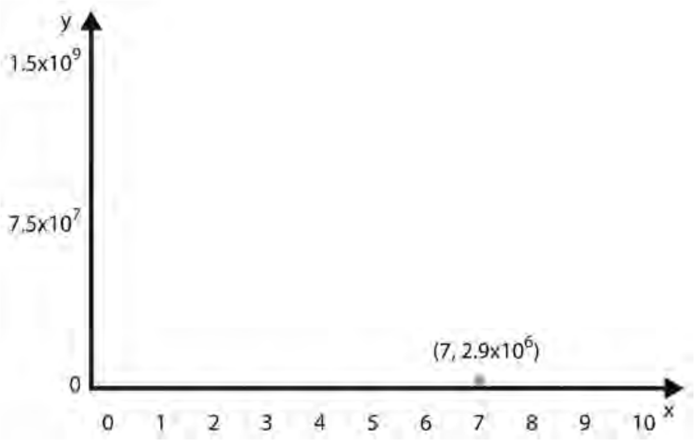
Next get your pH paper and dip those strips (a new one for each liquid) into each liquid. You can compare the color of the strip with a color code that comes with the pH paper to determine a number for the pH—something between 0 and 14.

Table 4.1

x value	*y* value
1.0	0.00000002
2.0	0.00006
3.0	0.004
4.0	0.8
5.0	25.0
6.0	3200
7.0	2900000
8.0	1400000000

Now for something completely different. Listed in Table 4.1 are a bunch of ordered pairs of numbers. You're going to try to graph these ordered pairs, with the *x* value going on the *x*-axis and the *y* value going on the *y*-axis. Figure 4.5 shows a sample graph with the first ordered pair plotted on the graph. Use the sample graph or draw one of your own to try to get all of the ordered pairs on one graph. Hint: It's pretty much impossible. You can give up once you know *why* it's impossible.

Figure 4.5



---

[5] Refer to Chapter 2 of this book for a definition of molarity. We use a capital M for molarity, so a label of 0.5 M means a chemical has a molarity of 0.5 moles per liter. Also, try to keep to low molarities of these substances.

Now use a calculator to take the logarithm of each of the *y* values and create a new table. I've done the first few and the last one for you below in Table 4.2.

Try graphing this second set of ordered pairs. Shouldn't be too difficult as long as you allow for negative *y* values.

Table 4.2

x value	Logarithm of *y* value
1.0	−7.7
2.0	−4.2
3.0	−2.4
4.0	?
5.0	?
6.0	?
7.0	?
8.0	9.1

More science stuff

You no doubt have used the terms *acid* and *acidic* in your everyday life. Many fruits contain citric acid, and we often refer to things we eat as being acidic or non-acidic. Movies let us know that acids can burn you (think *RoboCop*), especially the famous acids hydrochloric acid and sulfuric acid. People can have an acid tongue, but that's something different. Less common in everyday language are the words *base* and *basic*, at least as they apply to chemicals. The term *pH* is also common. We buy pH-balanced shampoos and check the pH of hot tubs and pools.

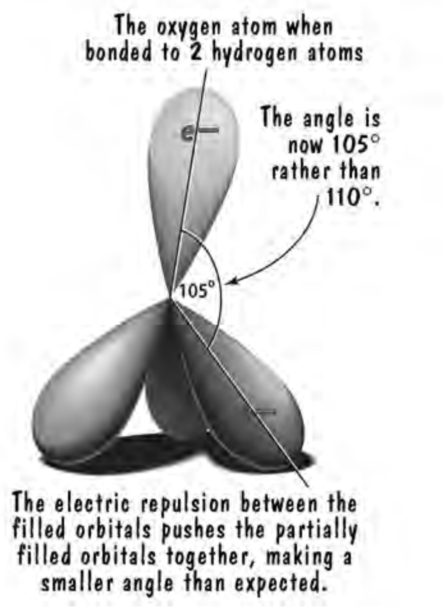
You probably knew before you even did the previous section that litmus paper helps determine whether a substance is an **acid** or a **base**. If blue litmus paper turns red when dipped in a liquid, the liquid is an acid. If red litmus paper turns blue in a liquid, the liquid is a base. If neither the red nor the blue litmus paper changes color, the liquid is neutral. pH paper takes this a step further and shows *how* acidic or basic a substance is.

Topic: Acids and Bases
Go to: *www.scilinks.org*
Code: MCB011

Great. Now if we only knew what acids and bases *are*! There are a number of ways to define acids and bases, and I'll give you two of them here. Before doing that, I need to remind you of how water interacts with other chemicals. Water molecules are **polar**, with one end being positive and the other negative. See Figure 4.6.

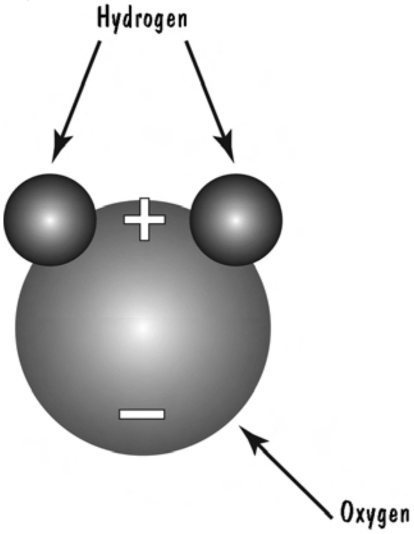
Figure 4.6 depicts the atoms hydrogen and oxygen. Be careful you don't confuse this drawing with the drawings in Chapter 3 that showed electron orbitals and how they interact. Of course, we can use the orbital drawings to explain exactly why oxygen is a polar molecule. Figure 3.29 is shown again on the next page. Notice that the two orbitals that do not bond with hydrogen atoms are filled with two electrons each. This makes the side of the molecule away from the hydrogen atoms more negative. Thus, a polar molecule.

Figure 3.29



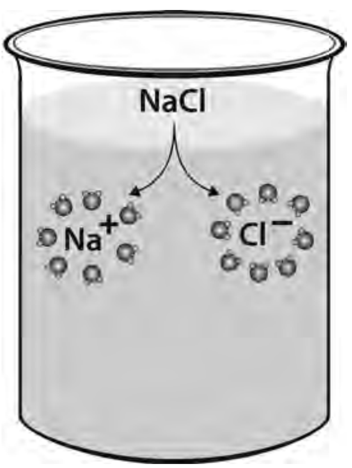
Because water molecules are polar, they interact electrically (you know, like charges repel and unlike charges attract) with other molecules. For example, when you put sodium chloride (table salt) in water, the attraction between different parts of the water molecule and the ions that make up sodium chloride are strong enough that water molecules separate the sodium chloride into positive sodium ions and negative chloride ions, as in Figure 4.7. Notice in this figure that the plus side of the water molecule (the hydrogen side) is attracted to the negatively charged chloride ion, and the negative side of the water molecule (the oxygen side) is attracted to the positively charged sodium ion.[6]

Figure 4.6



This tendency of water molecules to break apart other molecules is part of the first definition of acids and bases, known as the **Arrhenius theory of acids and bases** and named after (surprise!) a chemist named Svante Arrhenius. According to Arrhenius, an acid is a substance that increases the concentration of H+ ions in an aqueous (water) solution. So, hydrogen chloride (HCl) dissolved in water is an acid because the water breaks this molecule up into H^+ ions and Cl^- ions. Actually, this statement is a bit of a lie because H^+ ions (which are simply protons—take away an electron from hydrogen, which is com-

Figure 4.7



---

[6] When we break apart molecules with ionic bonds, we say they have been **dissociated**. From experience, you know that the salt disappears in this process because it has dissolved into the water. While molecules with covalent bonds can dissolve in liquid (such as sugar in water), covalent bonds don't dissociate in the process. They retain their molecular identity.

posed of one proton and one electron in its most common state, and all you're left with is a proton) don't exist by themselves in water. Instead, they latch onto water molecules to form what are known as hydronium ions, or H_3O^+. So really, an acid is anything that increases the hydronium ion concentration in an aqueous solution. Check out Figure 4.8.

Figure 4.8

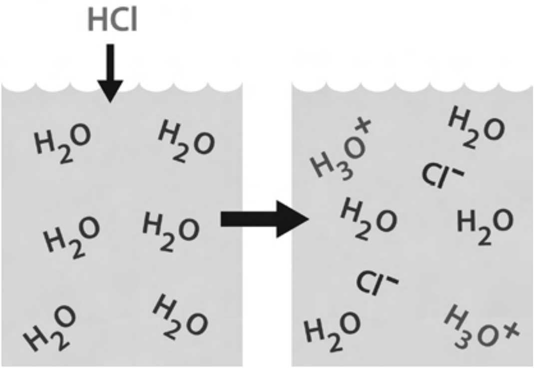
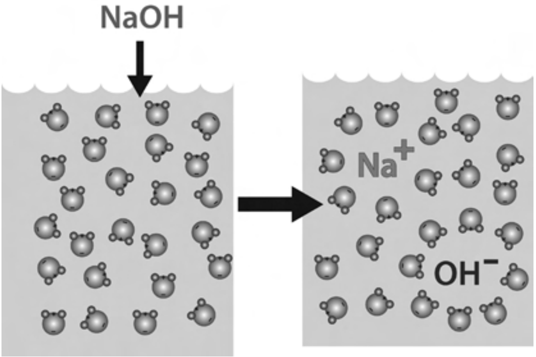


Figure 4.9



Arrhenius defines a base as anything that increases the concentration of hydroxide ions (OH^-) in an aqueous solution. For example, sodium hydroxide (NaOH) is a base because when you put it in water, the polar water molecules separate it into sodium ions (Na^+) and hydroxide ions (OH^-). Figure 4.9 shows this.

Okay, so what? What's so special about substances that form hydronium ions or hydroxide ions in water? Well, it has to do with what those substances do when they encounter other substances. If a substance produces lots of positively charged hydronium ions, then those ions have a tendency to interact with negatively charged ions. And if a substance produces lots of negatively charged hydroxide ions, then those ions tend to interact with positively charged ions. In other words, we define substances as acids and bases because that helps us figure out how they'll interact in chemical reactions, which is what most of chemistry is about.

I said I would give you two definitions of acids and bases, so I guess I'd better provide the second one. It turns out that molecules containing the hydroxide ion aren't the only ones that increase the concentration of hydroxide ions when dissolved in water. For example, here's what happens when ammonia (NH_3) dissolves in water:

$$NH_3\ (aq) + H_2O\ (l) \rightarrow NH_4^+\ (aq) + OH^-\ (aq)$$ [7]

---

[7] In case you forgot, (g) means the substance is a gas, (s) means the substance is a solid, and (aq) means the substance is in a solution. One classification that isn't represented here is (l), which indicates a substance that is itself a liquid.

NH_3 contains no hydroxide ions, yet when you put it in water it increases the hydroxide ion concentration. To account for examples such as this, we have what is known as the **Brönsted-Lowry** (Johannes Brönsted and Thomas Lowry) **theory of acids and bases**. In this definition, an acid is any substance that donates protons in a reaction, and a base is any substance that accepts protons in a reaction. This is basically (no pun intended) the same as the Arrhenius definition of acids because if a substance readily donates protons in a reaction, it will also increase the hydronium ion concentration when placed in water. The definition of a base, though, is expanded to include any substance that readily accepts protons in a reaction, and that substance doesn't have to contain hydroxide ions. Let's continue with our above example. There, NH_3 contributed to the OH^- concentration when dissolved in water. But NH_3 is also a proton acceptor (which we have defined as a base) because it picks up a proton to become NH_4^+, as in

$$NH_3\ (aq) + H_3O^+\ (aq) \rightarrow NH_4^+\ (aq) + H_2O\ (l)$$

Let's head back to your use of pH paper and why I had you attempt to graph those sets of numbers. You probably know that pH is a measure of how acidic something is, but you might not completely understand the pH scale, which is what I'm going to explain. First, let's talk about the two graphs I had you attempt to create. The *y* values in the first table (Table 4.1) have a wide range, from really small (0.00000002) to really large (1,400,000,000). If you make the graph so you can distinguish the small values, then the larger ones are on top of one another, and vice versa. Logarithms to the rescue. If you plot the logarithms of the *y* values instead of the original *y* values, then you can graph them easily.

Topic: pH Scale
Go to: *www.scilinks.org*
Code: MCB012

The concentrations of hydronium ions in aqueous solutions are like the numbers in Table 4.1 in that they cover a wide range of values. For example, a strong acid might have a concentration of hydronium ions of 0.01 M, and a weak acid might have a hydronium concentration of 0.000001 M.[8] We use logarithms to compare these numbers and graph them side by side. This leads us to the definition of pH, which is expressed in terms of hydrogen ion concentration, even though we realize that in the real world we're dealing with hydronium ions rather than hydrogen ions.[9] Here's the definition:

pH = negative of the log of the hydrogen ion concentration

---

[8] Remember that M stands for molarity, which is the number of moles of a substance per liter of solution.

[9] Many things in science are done in the name of tradition. The definition of pH was developed before chemists realized that hydrogen ions don't exist when in water. You will sometimes see the expression p(Hydronium) instead of pH, but for the most part, the term pH persists.

Chemists use brackets to indicate concentrations, so we rewrite the hydrogen ion concentration as $[H^+]$, and we have

$$\mathrm{pH} = -\log [H^+]$$

The reason for the negative sign is that most concentrations are less than 1.0, and as you saw previously, taking the log of these numbers gives you a negative number. So, the negative sign is just to make most pH values positive.

The nature of logarithms. How do logarithms (or logs) reduce a wide range of values to a small range of values? To understand that, you simply need to know the definition of a logarithm. To take the logarithm of a number, you ask, "10 raised to what power (exponent) will give me this number?" For example, log 100 is equal to 2 because 10 raised to the second power (10 squared) equals 100. Similarly, log 10,000 is equal to 4 because 10 raised to the fourth power equals 10,000. For numbers less than 1, negative exponents come into play. Log 0.01 is equal to −2 because 10 raised to the −2 power equals 0.01. So, just using the numbers I've mentioned here, the numbers 0.01, 100, and 10,000 become, after taking the logarithms of them, −2, 2, and 4. Voilà! A wide range of numbers reduced to a small range of numbers that are a bit easier to work with.

Assuming you used pH paper when I asked you to, you discovered the pH values of a few common liquids. The pH of lemon juice is about 2.3, the pH of vinegar is about 3.0, the pH of orange juice is about 3.5, and the pH of ammonia is about 11.0. Keep in mind that the higher the pH, the smaller the hydronium ion concentration and the greater the hydroxide ion concentration. The usual range of pH values chemists use is from 0.0 (highly acidic) to 14.0 (highly basic), even though values lower (and even negative) and higher are possible. A pH of 7.0 is considered neutral. And remember that pH is a logarithmic scale. That means that a pH of 5 is 10 times more acidic than a pH of 6 and 100 times more acidic than a pH of 7.

Before going on, I want to make a distinction between *strong* acids and bases and *concentrated* acids and bases. Hydrochloric acid is a strong acid because when it's placed in water, its separation into H_3O^+ ions and Cl^- ions is pretty much complete. Substances that readily donate protons are strong acids. However, you could have a solution of hydrochloric acid that has very little HCl dissolved in water or a lot of HCl dissolved in water. In other words, you could have a low *concentration* of a *strong* acid that isn't all that dangerous. If you'll recall, we measure concentration in molarity, so a 0.001 M solution of hydrochloric acid (0.001 moles of HCl per liter of solution) is much weaker overall (and would have a higher pH) than a 0.1 M solution of hydrochloric acid (0.1 moles of HCl per liter of solution), even though you would consider them both to be "strong" acids. And what I just said about acids applies to bases. The moral is to pay just as much attention to the molarity of an acid or base as you do to whether the acid or base is considered to be strong or weak.

Let's try an analogy to illustrate the difference between strong and weak on one hand and concentrated or not concentrated on the other hand. Suppose someone is going to set off firecrackers in your vicinity, and you want to be as safe as possible so as not to incur physical damage. There are tiny firecrackers and really big firecrackers. All other things being equal, you might choose to be in the vicinity of tiny firecrackers. That would be analogous to choosing a weak acid over a strong acid. But let's say this person is going to set off 100 fireworks, and there are two options—they'll be set off in a 3 × 3 meter closed room or in a 10 × 10 kilometer field. The different locations correspond to different concentrations. 100 firecrackers in a small room versus 100 firecrackers in a huge open field. The first option is a high concentration of firecrackers and the second is a low concentration of firecrackers. The difference in concentration is big enough that you could choose the large firecrackers (strong acid) in the large field (low concentration) and be safer than if you chose the small firecrackers (weak acid) in the small room (high concentration). So, the strength of an acid or base is important, but so is the concentration.

"That's correct!! You want to be near the **small** firecrackers. But you're going to have to move faster, Professor, if you want to make it to the next round."

"Our contestant is signaling he wants to be enclosed in the 3 m by 3 m case and **not** the 10 km by 10 km field. My, he certainly is intrepid, isn't he."

Even more things to do before you read even more science stuff

We covered what acids and bases are, so it's time for reactions. For this activity, you'll need some calcium hydroxide (about 0.1 M concentration) dissolved in distilled water (get the powder from a science chemical supplier)[10], a straw, a small amount of *clear* carbonated soda (club soda or lemon-lime soda), a pair of safety goggles, and your breath. Pour a small amount of the calcium hydroxide solution (the liquid you made, not the powder) in a clear glass and notice what color it is (clear, yes?). Use the straw to blow bubbles into the liquid (possible splashback is why you have goggles). Notice any change in the calcium hydroxide (there *should* be a change). Rinse out the glass and pour a small amount of clear carbonated soda into it. By definition, this is clear. Then pour a small amount of calcium hydroxide into the glass and notice any change.

Even more science stuff

If all went well, cloudiness happened. When you blow bubbles into clear calcium hydroxide, it turns cloudy. When you pour clear calcium hydroxide into clear carbonated soda, the solution turns cloudy. The explanation for both of these is that you observed an acid-base reaction. In each case, the cloudiness came from the interaction of an acid—carbonic acid (H_2CO_3)—and a base—calcium hydroxide. A carbonated soda already has carbonic acid in it. When you blow bubbles into calcium hydroxide, the carbon dioxide (CO_2) from your breath interacts with the water in the calcium hydroxide solution to form H_2CO_3. Here's the balanced equation for the interaction of carbonic acid and calcium hydroxide:

$$\text{carbonic acid} + \text{calcium hydroxide} \rightarrow \text{calcium carbonate} + \text{water}$$

$$H_2CO_3\text{ (aq)} + Ca(OH)_2\text{ (aq)} \rightarrow CaCO_3\text{ (s)} + 2H_2O\text{ (l)}$$

Quick review. How did I know the formulas for carbonic acid, calcium hydroxide, and calcium carbonate? Yeah, I could have looked them up, but I didn't. I simply used the Periodic Table along with the knowledge that OH^- has a negative charge (an extra electron) and that CO_3^{2-} has a −2 charge (two extra electrons). Also, if you don't know how I balanced the equation by using two waters, it might be worth reviewing how to balance equations.

Calcium carbonate is an insoluble salt (doesn't dissolve in water), which is why things turned cloudy. In fact, all acid-base reactions have the same result.

---

[10] The concentration of your dissolved calcium hydroxide isn't critical. About a teaspoon of powder in a cup of distilled water will do the trick. You can also buy solutions of calcium hydroxide so you don't have to bother dissolving any powder into distilled water.

You add the acid and base together, and you get a salt (there are many kinds of salt that can be produced) plus water. The production of salt isn't always easy to see, though. For example, if you mix vinegar (an acid) and ammonia (a base), you won't see any cloudiness. The salt produced in this case (ammonium acetate) stays dissolved in the solution.

Now, you might be thinking that we could figure out this reaction without the whole notion of acids, bases, and acid-base reactions. You would be right. You could simply look at the relative attractions for electrons of the atoms and compounds involved and know that certain bonds are going to be stronger than others.[11] You could also know that carbonic acid is very likely to give up a proton and calcium hydroxide is very likely to accept a proton in this reaction, without classifying them as acids and bases. The bottom line, though, is that there are so many acid-base reactions that it's useful for them to have a category all by themselves. Simply knowing that you are starting with acids and bases tells you what will happen in the reaction without going through all that other stuff.

And even more things to do before you read even more science stuff

You'll need a few special chemicals for this one. The first is methylene blue, which is an acid-base indicator. You just need a tiny amount (maybe 10 drops) of this. The second is glucose (also known as dextrose anhydride), which you can get as a powder in small amounts from a science supply outlet. The third is potassium hydroxide, which you can get as a powder or solution. You'll also need a large glass container with a tight-fitting lid. To make it look all chemistry-like, people traditionally use a 500 ml Erlenmeyer flask with a rubber stopper (see Figure 4.10), but that's not absolutely necessary. Just make sure there's no danger of spilling what you end up with because it can irritate your skin.

Figure 4.10

Now, if I wanted to be mean, I'd give you the molarities of the solutions you need and have you use what you learned in Chapter 2 to make up the solutions. Because I'm a nice guy, though, here's what to do.

---

[11] How we do this is discussed in the first *Stop Faking It!* chemistry book. It involves the concept of electronegativity (a measure of affinity of atoms for electrons) and comparison of the numbers associated with electronegativity.

- Do the following preparations within ten or fifteen minutes of doing the activity. Leave things around for a day, and it won't work.
- Add about 8 grams of potassium hydroxide (KOH) to about 300 ml (that's milliliters) of distilled water in the flask and swirl it to mix thoroughly.
- Add about 10 grams of glucose (dextrose) to this solution and allow it to dissolve.
- Add 6 to 8 drops of methylene blue to the flask and swirl things around again.
- Let everything rest until the solution becomes clear.

Now for the fun part. Stopper the flask and shake up the solution inside. It should turn blue. Let it sit for a bit and it should clear up. Shake it again, and it should turn blue again. You can repeat quite a few times before there's no longer a color change.

And even more science stuff

First, the quick and dirty explanation of what happened. When you shake up the solution, you dissolve oxygen from the air in the container into the solution. The oxygen combines with the molecules in the solution, resulting in a change to the blue color. Then there's a spontaneous reaction in which oxygen is released back into the air, and the solution changes to clear.

Now, that's a pretty general explanation, so obviously we need more detail. There are a fair number of chemicals involved, so I'll just focus on the important ones. When you first mix things together, a couple of things happen. The first is listed below.

$$\text{glucose} + \text{hydroxide ions} \rightarrow \text{gluconic acid} + \text{water} + \text{electrons}$$

Using chemical formulas, this equation looks like

$$HOCH_2(CHOH)_4CHO\ (aq) + 3OH^-\ (aq) \rightarrow HOCH_2(CHOH)_4CO_2\ (aq) + 2H_2O\ (l) + 2e^-$$

Don't get flustered by how complicated the molecules are. We're just going to look at certain parts of the equation to understand what happens to those parts. The hydroxide ions in this equation come from the potassium hydroxide you added to the solution. Also, note that e^- is just the symbol for an electron. Let's look at just the glucose and the gluconic acid molecules. I'll write them to emphasize what's happening at the far right side of each.

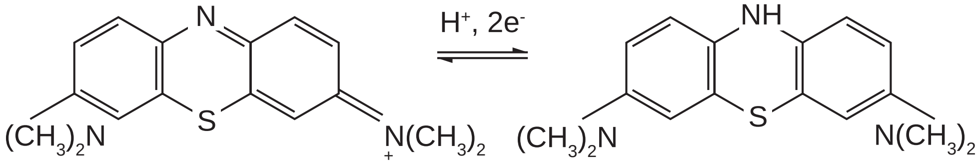
$$\underset{\text{glucose}}{HOCH_2(CHOH)_4C\text{-}OH} \qquad \text{and} \qquad \underset{\text{gluconic acid}}{HOCH_2(CHOH)_4C\text{-}O_2}$$

The only difference between these two molecules is that glucose has an OH on the end and gluconic acid has an O_2 on the end. We know that OH^- ions have an overall charge of −1, so, because the entire molecule is neutral, we can think of the rest of the molecule as having a +1 charge. For gluconic acid, we need to look at the O_2. Oxygen atoms need two electrons to fill their outer shells (refer to the Periodic Table), so when oxygen combines with other molecules, they share two electrons and those shared electrons spend more time near the oxygen.[12] Because there are two oxygen atoms (O_2) instead of one, there are four electrons (two per oxygen atom) spending more time near the O_2. *In a sense,* then, the O_2 at the end of gluconic acid has a −4 charge. Because this molecule is neutral, that means that *in a sense*, the rest of the molecule has a +4 charge. I emphasize the words *in a sense* because we're not talking about an ionic bond, in which electrons have jumped from one atom to another. We're just focusing on where the electrons spend more of their time and assigning plus and minus charges based on this observation.[13] Anyway, what's happened is our glucose molecule, in becoming a gluconic acid molecule, has gone from a +1 "charge" to a +4 "charge." Because the molecule has become more positive, it's as if the molecule has *lost electrons* (losing negative electrons increases the positive charge). When a molecule loses electrons, we say it has undergone **oxidation**. So we would say that the glucose has been oxidized. In this case, the molecule actually added an oxygen molecule, and that's what historically led to this kind of reaction being called an oxidation, but oxidation can take place without oxygen being involved.

Now on to the next equation, and remember that we're just looking at what happens when you first mix the chemicals.

methylene blue (blue form) + glucose → methylene blue (clear form)
+ a bunch of other chemicals

Now let's just focus on what happens to the methylene blue and represent it as the molecular structure rather than just formulas. Again, don't get disturbed by the complexity of the molecule. We'll only focus on small parts of it.



---

[12] Recall that this happens in water (H_2O) molecules. The shared electrons spend more time around the oxygen atom, making the molecule polar.

[13] The "charges" I'm talking about here are referred to in chemistry circles as **oxidation states**. Just thought I'd mention it in case you come across the term somewhere.

That two-way arrow simply means that this reaction is reversible—it can occur in either direction. In comparing the two molecules, you can see that the one on the left has an N at the top center, while the one on the right has an NH. Also, the plus sign on the lower right of the left molecule has disappeared on the right molecule. The symbols above the arrows indicate how this happened. We have added an H^+ ion and also added two electrons to the molecule on the left. These come from the glucose molecule in the word equation I wrote first. One of the added electrons makes the NH at the top neutral (remember, we added an H^+ ion), and the other gets rid of the plus sign at the lower right. Because we added two negatively charged electrons and only one positively charged H^+ ion, the molecule has gained one electron overall. When a molecule gains an electron, we refer to that as a **reduction**.

Topic: Redox Reactions
Go to: *www.scilinks.org*
Code: MCB013

The reaction I have just described is an example of a special class of chemical reactions known as **oxidation-reduction** reactions, or **redox** reactions for short. When an atom or molecule *loses* electrons in a reaction, we say the substance is oxidized. When an atom or molecule *gains* electrons in a reaction, we say the substance is reduced.[14] In a redox reaction, at least one atom or molecule gains electrons, and at least one atom or molecule loses electrons.

What happens when a substance is reduced or oxidized can be confusing, but there's an easy way to remember it. Just think of a lion, and remember that **LEO** says **GER**. This stands for **L**oss of **E**lectrons is **O**xidation and **G**ain of **E**lectrons is **R**eduction. And no, I didn't come up with that. I'm not that clever.

As with our previous classifications of reactions, classifying reactions as oxidation-reduction is simply a tool to help our understanding. In that sense, there's nothing new at all in this chapter. Chemicals get together and do their thing. If we can place groups of reactions in separate classifications, however, we can better predict what might happen in other circumstances. That said, I should let you know that there are still many other classifications of reactions that chemists use. The three reactions in this chapter aren't the only ones, but they are major ones and illustrate the classification process. And if you think I've done something wrong by not including all the different types of reactions, I ask you to remember the purpose of this book, which is to help you develop a deep understanding of basic concepts. Don't think of this as a comprehensive resource for all your chemistry needs.

---

[14] As I stated earlier, the reason for the term oxidation-reduction has to do with the fact that these kinds of reactions were at first thought to involve only the combining of oxygen with other molecules, as happens in this reaction. But this class of reactions is much more general and doesn't necessarily include oxygen.

Chapter summary

- A combustion reaction is one in which oxygen combines with a hydrocarbon to produce carbon dioxide and water.
- A catalyst is a substance that speeds up or improves the efficiency of a chemical reaction. The catalyst itself is unchanged as a result of the reaction.
- Activation energy is the energy input required to get a chemical reaction to proceed in a certain direction. A catalyst effectively lowers the activation energy for a reaction.
- An acid is any chemical that, when dissolved in water, increases the concentration of hydronium ions. Alternatively, an acid is anything that donates protons in a reaction.
- A base is any chemical that, when dissolved in water, increases the concentration of hydroxide ions. Alternatively, a base is anything that accepts protons in a reaction.
- pH is a logarithmic scale that measures the concentration of hydronium ions in a solution. The scale ranges from 0 (highly acidic) to 14 (highly basic).
- There is a difference between the strength of an acid or base and the concentration of an acid or base. The strength describes how completely the acid or base donates or accepts protons. The concentration tells how much of an acid or base a solution contains.
- Combining acids and bases always results in the production of a salt and water.
- When an atom or molecule loses electrons during a reaction, we say it is oxidized. When an atom or molecule gains electrons during a reaction, we say it is reduced.
- Oxidation-reduction (redox) reactions involve at least one substance being oxidized and at least one substance being reduced. Keeping track of the oxidation and reduction is a useful way to analyze what happens in a chemical reaction.

Applications

1. Most of us use a catalyst reaction almost every day. Cars have **catalytic converters** that help remove harmful substances before they escape out the exhaust pipe. The reason engines produce harmful substances is that the combustion, or burning, in the engine is never complete. After the engine does its thing, the following harmful substances remain: nitrogen oxide,

nitric oxide, carbon monoxide, and unburned hydrocarbons.[15] To get rid of these bad substances, we'd like the following reactions to occur:

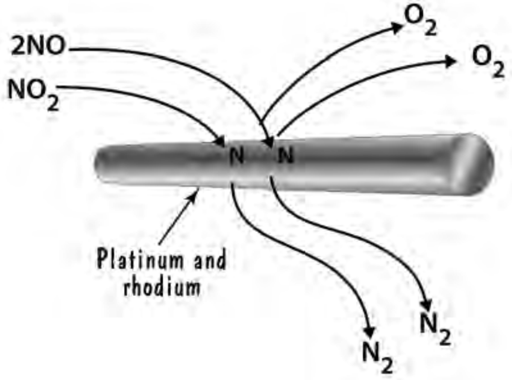
nitrogen oxide → nitrogen + oxygen
nitric oxide → nitrogen + oxygen
carbon monoxide + oxygen → carbon dioxide
hydrocarbons + oxygen → carbon dioxide + water

The balanced chemical equations for these reactions are below. Note that C_xH_y is the generic term for hydrocarbons that can contain any number of carbon and hydrogen atoms (hence the *x* and *y* subscripts rather than numbers). Because we don't have exact numbers for *x* and *y*, the last equation isn't balanced.

$$2NO_2\ (g) \rightarrow N_2\ (g) + 2O_2\ (g)$$
$$2NO\ (g) \rightarrow N_2\ (g) + O_2\ (g)$$
$$2CO\ (g) + O_2\ (g) \rightarrow 2CO_2\ (g)$$
$$C_xH_y\ (s) + O_2\ (g) \rightarrow CO_2\ (g) + H_2O\ (g)$$

Okay. These are the reactions we want to happen, and they do happen somewhat as the byproducts of the engine leave the exhaust. The problem is that they don't happen all that quickly, leaving bad stuff heading out into the atmosphere. That's why we use catalysts that lower the activation energies for these reactions and help them proceed at a rapid rate. A catalytic converter is inserted in the exhaust system of a car so that it receives the bad stuff after it's left the engine. The catalytic converter contains a ceramic core coated with very expensive metals—platinum, rhodium, and palladium—and has two sections. The first section deals with nitrogen oxide and nitric oxide. This first part has platinum and rhodium covering the ceramic. These metals rip nitrogen atoms off the nitrogen oxide and nitric oxide. The leftover oxygen atoms form oxygen molecules, and the nitrogen atoms held by the metals combine to form harmless nitrogen. Figure 4.11 illustrates the process.

Figure 4.11

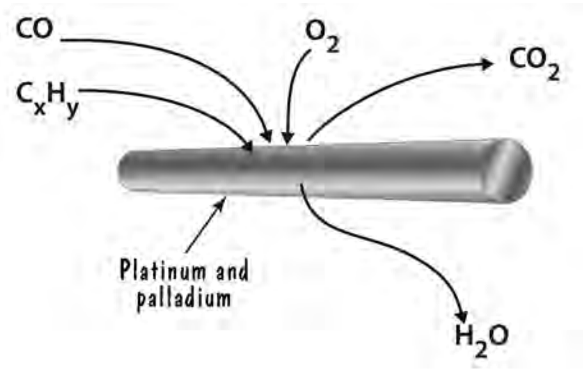


---

[15] Hydrocarbons are any molecules that contain carbon and hydrogen.

The second part of the catalytic converter deals with the carbon monoxide and the unburned hydrocarbons. Here the ceramic is coated with platinum and palladium. These metals attract carbon monoxide, hydrocarbons, and oxygen. By trapping these molecules, the metals make it easier for them to get together in a chemical reaction. Figure 4.12 illustrates this.

Figure 4.12



As with all catalyst reactions, the catalysts (the expensive metals) help make it easier for a chemical reaction to take place while not changing themselves in the process. The metals get the other atoms together, or rip molecules apart, while staying on the ceramic. This is a good thing because you want a catalytic converter to last a while.

2. The reactions that take place in the two parts of a catalytic converter are referred to as the *reduction* reaction and the *oxidation* reaction. Can you figure out which is which? Aside from the fact that carbon atoms latch onto oxygen atoms, which indicates an oxidation, you can figure this out by looking at the "charge" (which we call an oxidation state) on the nitrogen atoms and the carbon atoms in each reaction. First, consider the NO molecule. There is a double bond between the atoms, and the electrons in those bonds spend more time around the oxygen than they do around the nitrogen. This is because oxygen has a greater affinity for electrons than nitrogen. This means that we can think of the nitrogen in NO as having a slightly positive charge or oxidation state. As with our earlier examples of oxidation and reduction, it's not that the nitrogen is positively charged and the oxygen negatively charged, but in the molecule they sort of act that way. One product of the reaction is N_2. In this molecule, each nitrogen atom is bonded to another nitrogen atom, so the electrons in the bonds don't prefer one atom over another. Thus, in N_2, the "charge" on each nitrogen is zero. So, nitrogen has effectively *gained* electrons in the reaction (going from a plus charge to zero charge). This means that nitrogen is *reduced,* and the reactions that take place in the first part of a catalytic converter are known as *reduction reactions.* Note, though, that in labeling these as reduction reactions, we're only focusing on what happens to the nitrogen atoms. While the nitrogen is being reduced, the oxygen is being oxidized—you can't have a reduction of one substance without a corresponding oxidation of another substance.

Okay, what about the second part of the catalytic converter? Once again, we just have to look at one of the reactions, such as hydrocarbons adding to oxygen to produce carbon dioxide and water. Let's just look at the carbon atoms. In a hydrocarbon, the carbon atoms are bonded to one another or to hydrogen atoms. When carbon and hydrogen atoms get together, the different affinities for electrons dictate that the electrons will be more attracted to the carbon atoms, so the carbon atoms will have a negative "charge" in hydrocarbons. After the reaction, carbon atoms are bound to oxygen atoms (in CO_2). Again, affinities for electrons tell us what will happen. Electrons spend more time around the oxygen atoms, so the carbon atoms in CO_2 have a positive "charge." Carbon goes from a negative "charge" to a positive "charge."[16] This process involves a loss of electrons, so carbon is *oxidized* in this process. Therefore, the reactions that take place in the second part of the catalytic converter are called *oxidation reactions*.

3. Molecular formulas for acids and bases often are written so it's obvious that they are acids or bases. For acids, the hydrogen ion (proton) that the molecule readily loses is placed at the beginning of the formula rather than lumped with the rest of the hydrogen atoms in the molecule. For example, acetic acid (vinegar) is $HC_2H_3O_2$, with that first H (or sometimes more than one H) being the one that jumps off easily. Lactic acid is $HC_3H_5O_3$, and citric acid is $H_3C_6H_5O_7$. Two structures that easily accept protons, and thus make certain molecules bases, are NH and NH_2. For easy identification, many bases are written with these structures tacked onto the end of the molecular formula. For example, we write the bases dimethylamine, methylamine, and aniline as $(CH_3)_2NH$, CH_3NH_2, and $C_6H_5NH_2$.

4. Acids, bases, and acid-base reactions play a big part in the functioning of the human body. For one thing, your body is pretty sensitive to changes in pH. Anything too acidic or too basic will generally be a problem. Even though your stomach relies on acids to break down foods, too much acid in your stomach can hurt. To counteract this, you take a base called an antacid (duh!). Because bases snarf up hydronium ions, antacids make your stomach less acidic (higher pH). Of course, you have to be careful about what bases you put in your stomach. Ammonia is a pretty strong household base, but it's much too strong to put in your body. When my father had an upset stomach, he always swallowed baking soda and water—a base that is quite a bit stronger than your usual antacid. That usually solved his problem quickly but violently. I don't recommend it!

---

[16] Throughout this application, I have put the word *charge* in quotation marks. To be absolutely correct, I should use the term *oxidation state*. I just want to say once again that these atoms and molecules carry a charge only *in a sense* because while the atoms and molecules are connected, the shared electrons spend more time around one than the other.

5. Sometimes chemical equations are difficult to balance just by trying to balance the number of each kind of atom on both sides of the equation. If the equation is an oxidation-reduction reaction, there's a surefire method for balancing the equation. I won't give a detailed example, but I can tell you that the method involves looking at the oxidation states (that "charge" that atoms and molecules have, sort of, when combined) of the atoms and molecules and balancing the electrons lost with the electrons gained. It's a nifty trick that works every time, so if you find yourself with chemical equations that you just can't balance by inspection, you might want to look up the method and learn it.

5 Chapter

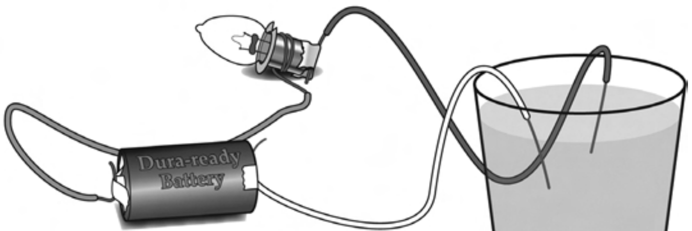
Electro-Luminescence

The title of this chapter is a compact way of saying that I'm going to cover two separate branches of chemistry. The first is electrochemistry and the second is the interaction of light with matter. I suppose I could spend an entire chapter on each of these areas, but that would work against the purpose of these books. I want to increase your understanding so that you can tackle regular textbooks, not create a new textbook.

Things to do before you read the science stuff

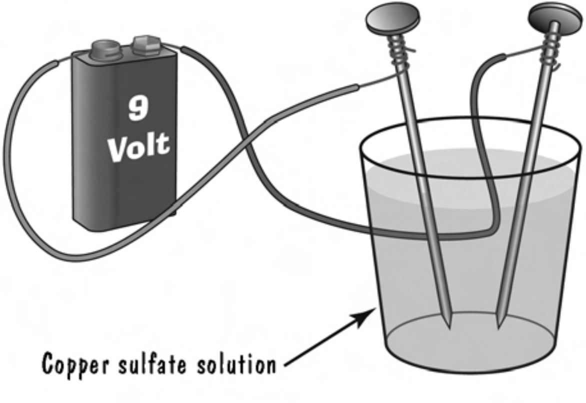
The first thing I'm going to have you do is a repeat of an activity in the first *Stop Faking It!* book on chemistry, so if you've already done it, rely on that great memory of yours. You need a flashlight bulb, a 1.5-volt battery, three wires about 30 centimeters in length each (stripped at the ends if they're insulated), masking tape, a cup of water, and table salt. You can substitute a threaded bulb and base (get this at an electronics store and make sure its rating is around 1.5 to 3 volts) and wires with alligator clips on the end to make the whole thing a bit easier. Set everything up as shown in Figure 5.1. No salt is involved yet, and make sure the wire ends in the water don't touch each other.

Figure 5.1



If you've done everything correctly, nothing should happen. Now add a whole bunch of salt to the water and stir it up. With everything in place as in Figure 5.1, the lightbulb should now light. If not, check to make sure all your connections are secure and that the battery isn't dead.

Figure 5.2



Keep the battery,[1] wires, masking tape, and cup handy. Add to this list a couple of nails and a small amount (about a spoonful) of copper sulfate crystals (a special chemical—check out science supply stores). You can keep the light bulb in the circuit if you want, but it's not necessary. Strip more insulation off each of your wires so you can wrap a fair amount of this bare wire around the nails. Dissolve the copper sulfate crystals in a cup of water (stir until dissolved) and then set things up as in Figure 5.2.[2]

---

[1] What you're about to do will go a lot faster if you use a 9-volt battery instead of a 1.5-volt battery. The former also has the advantage that alligator clips hook easily on its terminals. Of course, if you're patient, then a 1.5-volt battery will do just fine.

[2] This will work better with distilled (not de-ionized) water, but that's not absolutely necessary.

Make sure the nails in the solution do not touch. When setting this up, connect all the components *before* placing the nails in the solution.

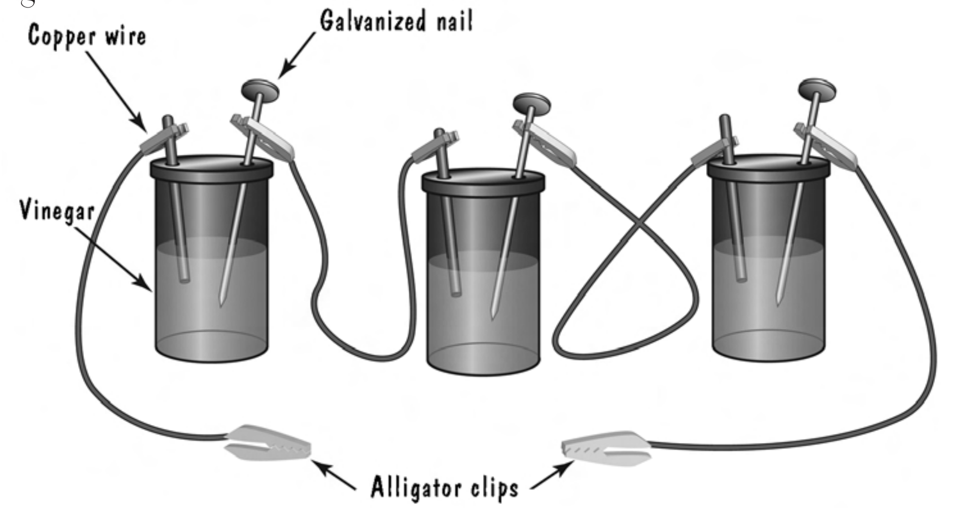
Wait a while, at least five or ten minutes. After this time, pull the nails out of the solution and notice any change. There should be a change. The longer you keep the nails in the solution, the greater the change.

One final electrical thing to do is create your own battery. This takes patience, but stick with it. It's rewarding once things work. Here's what you need:

- three galvanized nails (galvanized is important)
- three short (3–4 cm) strips of copper wire (thicker is better—about 16–18 gauge)
- four wires with insulation stripped from the ends (alligator clips are a big plus here)
- an LED (light-emitting diode—check RadioShack or other electronics store)[3]
- three cups or empty 35 mm film canisters with lids (You can get them free at most photo finishing places—the folks at Wally World [Walmart] are especially kind to teachers.)
- vinegar (about a cup)

Figure 5.3 shows how to set up the activity using film canisters. The lids on the film canisters are nice for securing the nails and copper wires so they don't

Figure 5.3



---

[3] For large quantities of LEDs, snoop around the internet for low prices. I was able to find them cheapest at *www.ledsupply.com,* but there's no guarantee that this site will be up and running by the time this book is in print.

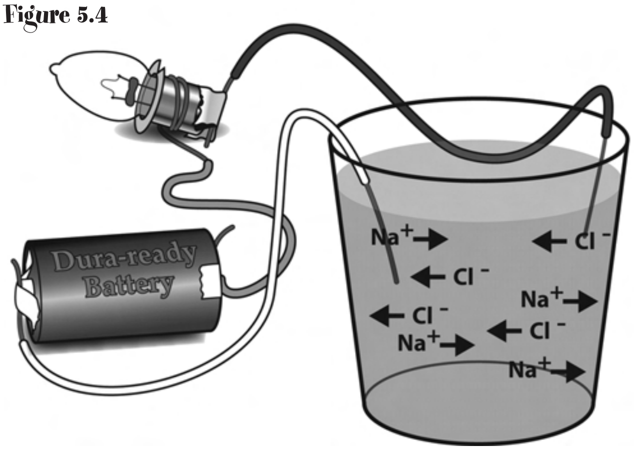
touch. Make sure you put enough vinegar in each canister so both the nail and the copper wire are submerged.

Before you hook up the LED, you need to know which connection is positive and which is negative. The shortest wire leading from the LED is the negative connection, and it's also the wire that's closest to the flat side of the plastic part of the LED. Once you have the positive-negative thing figured out, connect the free wire leading from the copper wire to the positive lead on the LED (the long wire) and connect the free wire leading from the zinc-covered nail to the negative lead on the LED (the short wire). You should see the LED light up. If not, check all the connections and try again. If you still don't get a lit bulb, keep connecting and reconnecting wires and make sure the nails and the copper are scraped clean where they connect to the wires or alligator clips. If you just can't get it to work, hook your LED up to the 1.5-volt battery (positive terminal of battery to long wire on the LED and negative terminal of battery to short wire on the LED) to make sure the LED is working. If so, then it's back to fiddling around with the connections. Before giving up completely, use the tried-and-true method of having a child work on it. They're better at this than we are.

The science stuff

Most people associate the word *electricity* with wires and the motion of electrons through those wires. Makes sense, because that's the kind of electricity we encounter on a day-to-day basis. In the previous section, though, all of your electrical circuits involved liquids—saltwater, copper sulfate dissolved in water, and vinegar. In each case, ions (charged atoms or molecules) in the liquids moved around instead of electrons. An electric current is defined as any movement of charges, not just the movement of electrons. Because we're dealing with ions in addition to electrons, we call the process **electrochemistry**.

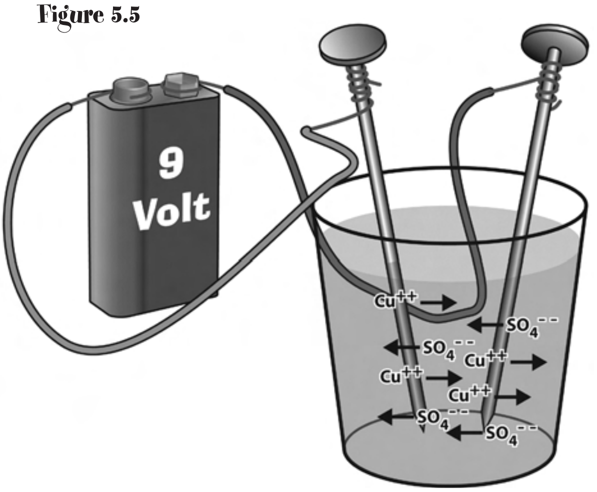
Figure 5.4



We've discussed before how salt (NaCl) separates into positive sodium ions (Na^+) and negative chloride ions (Cl^-) when you put it in water. When you hook up a battery to wires submerged in saltwater, the positive sodium ions are attracted to the negative terminal of the battery, and the negative chloride ions are attracted to the positive terminal of the battery. This movement of positive and negative ions in the water constitutes an electric current, and you have a complete circuit that lights the lightbulb. Check out Figure 5.4.

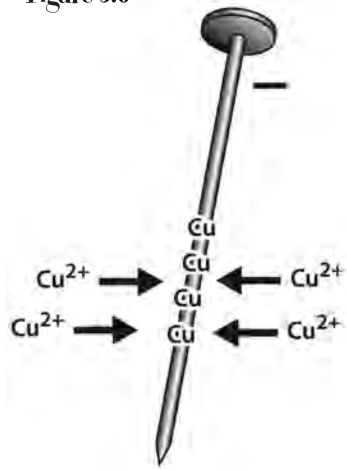
Switching to copper sulfate ($CuSO_4$) solution, you have a similar situation. When you dissolve copper sulfate crystals in water, the water molecules separate them into positive copper ions (Cu^{2+}) and negative sulfate ions (SO_4^{2-}). Even though the water molecules tend to isolate these ions, the ions are still free to move toward things that have the opposite charge. So, the Cu^{2+} ions gather around the wire that's connected to the negative terminal of the battery, and the SO_4^{2-} ions gather around the wire that's connected to the positive terminal of the battery. See Figure 5.5.

Figure 5.5



The interesting thing is what happens at the negative terminal. Cu^{2+} ions grab onto free electrons and become regular copper. Whatever metal you have at the negative terminal becomes coated with copper.[4] And yes, that sludge that forms around the negative terminal is also copper. See Figure 5.6.

Figure 5.6



The official name for this process is **electroplating**. Wonder no longer how people can create a ring plated with gold or a piece of coated silverware or a piece of silver-plated jewelry. It's all about creating ions of the plating material and using electric attraction to seduce those ions from one place to another. Of course, this doesn't always have to take place in a liquid. Often it's done within a vacuum (not much air), with molecules of one substance getting charged up and then being attracted to an object that has the opposite charge.

---

[4] If you ignored my directions and put the nails in the solution before the rest of the circuit was hooked up, you probably ended up with both nails being coated with copper. This is because simply placing a nail in copper sulfate solution will result in some copper ions attaching to the nail. If you don't submerge the nails until the circuit is ready, then copper ions are immediately repelled from the positive terminal when you place it in the water. Then few copper ions will attach to that terminal.

The vacuum process is actually much more efficient than the liquid process because there are fewer other molecules (such as deposited minerals in the liquid) to get in the way of the transfer of metal from one place to another.

The final thing you did in the previous section was make a battery. One film canister with vinegar isn't sufficient to light the LED, which is why I had you connect three of them together. One alone will produce enough voltage (about 1 volt) to light the bulb, but it won't produce enough electric current to light the bulb.[5] Three of these batteries together do the trick.

Now, to understand why our battery causes electrons to move in the outer circuit (the wires and the bulb), we have to look at the chemical reactions taking place with the zinc (this is the metal coating on galvanized nails), copper, and vinegar. Each metal reacts with the vinegar, which is acetic acid dissolved in water. Knowing what we know about what happens to acids in water, we can figure out that acetic acid ($HC_2H_3O_2$) will separate into ions, giving up a proton (a hydrogen nucleus). That free proton readily combines with water to form a hydronium ion (H_3O^+), leaving a negatively charged ion that we call acetate ($C_2H_3O_2^-$). Now we can look at what happens when these ions come in contact with the zinc that covers the galvanized nail.

$$\text{zinc} + \text{vinegar} \rightarrow ???$$

$$\text{zinc} + \text{hydronium ions} + \text{acetate ions} \rightarrow ???$$

In symbols, this is

$$Zn\ (s) + H_3O^+\ (aq) + C_2H_3O_2^-\ (aq) \rightarrow ???$$

So what goes on the right-hand side of this equation? Well, the extra H^+ in the hydronium ion isn't held to that ion strongly, so it can be fickle and hook up with other atoms or ions. Enter zinc, which would love to get rid of a couple of electrons to get a filled outer shell. The loosely held zinc electrons leave the zinc atom and combine with the H^+ from the hydronium ions. The remaining positively charged zinc ions remain in solution, as do the acetate ions. This leads to the following equation.

---

[5] I'm not going to get into electric circuits here, but I can maybe help just a bit with why one canister doesn't do the trick, but three canisters do. All batteries, including your vinegar batteries, have what is known as *internal resistance*. This internal resistance impedes the flow of electric current. The less internal resistance a battery has, the more current it can deliver to a device such as an LED. Your vinegar battery has a relatively large internal resistance, so even though one or two canisters have enough voltage to light the LED, they can't produce enough current to light it. You need three of them. A regular 1.5-volt battery has a small internal resistance, so just one will light the LED.

zinc + hydronium ions → zinc ions + hydrogen gas + water

$$Zn\ (s) + 2H_3O^+\ (aq) \rightarrow Zn^{2+}\ (aq) + H_2\ (g) + 2H_2O\ (l)$$

As far as our battery is concerned, the only thing to focus on is the fact that neutral zinc becomes a positively charged zinc ion while giving up a couple of electrons. That's easy to write as what's known as a **half-reaction**, and it looks like this:

$$Zn \rightarrow Zn^{2+} + 2e^-$$

where e^- represents an electron.

At the copper wire, a similar reaction takes place. In fact, all you have to do is replace Zn (zinc) with Cu (copper) in the previous equations, and you have what happens at the copper wire. Here's the chemical equation:

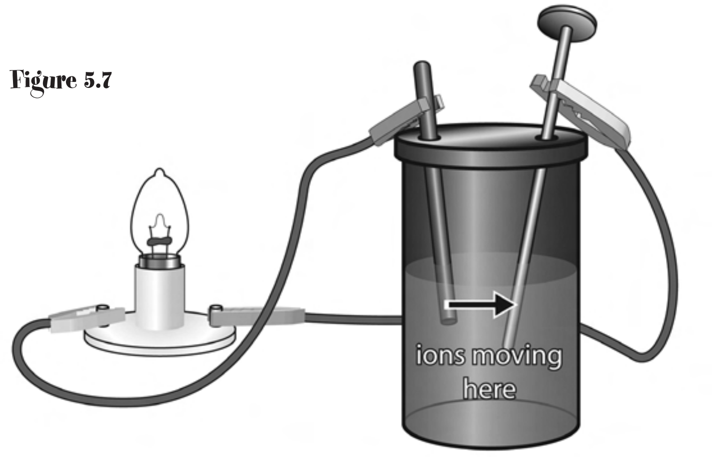
$$Cu\ (s) + 2H_3O^+\ (aq) \rightarrow Cu^{2+}\ (aq) + H_2\ (g) + 2H_2O\ (l)$$

And just focusing on the copper, we have

$$Cu \rightarrow Cu^{2+} + 2e^-$$

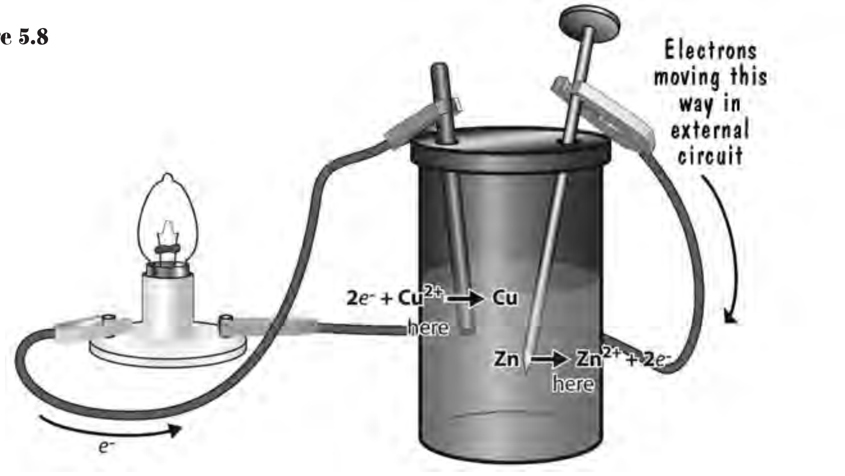
So, we're producing electrons at the zinc-covered nail and we're producing electrons at the copper wire. How in the world does that make a battery? Well, let's hook up an outside circuit to this setup, as in Figure 5.7.

Figure 5.7



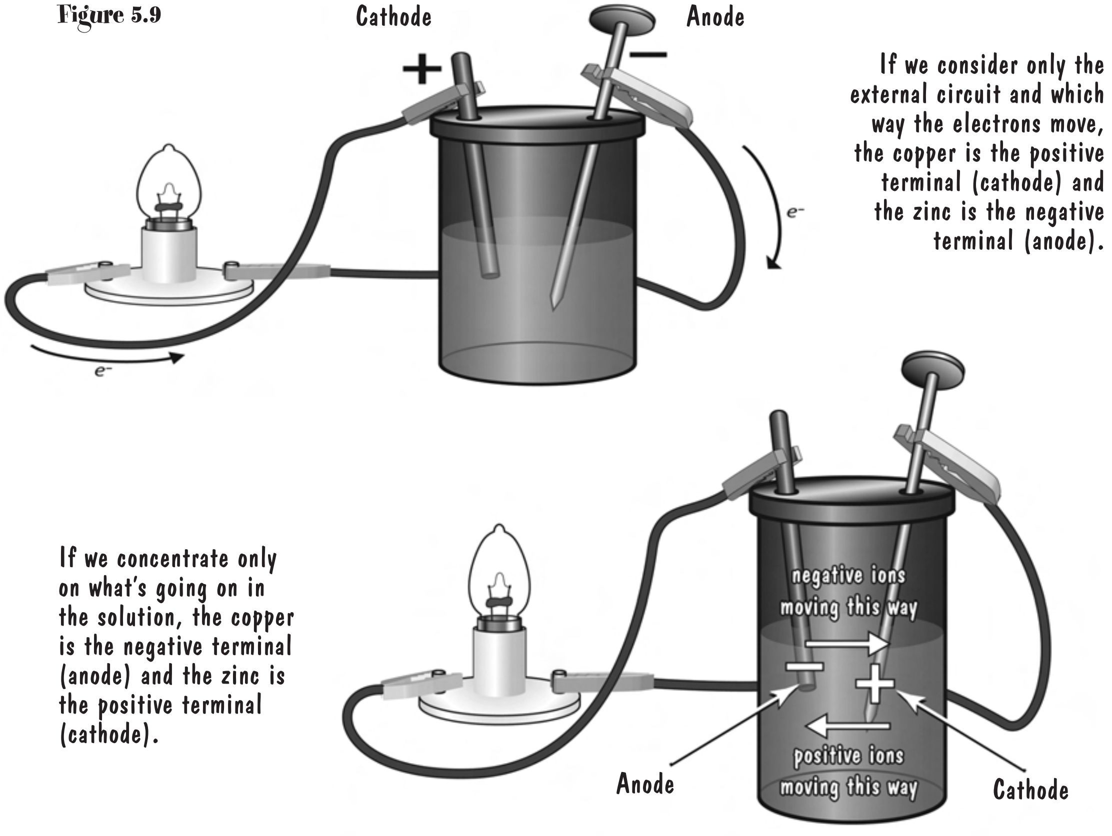
Now we have a complete circuit. Electrons can carry electric current in the wire, and ions (zinc and acetate) can carry electric current in the liquid. With everything connected, there is communication between what's happening at the nail and what's happening at the copper wire. Any change at one place will transmit electrically to the other place. And here is where, as is often the case in chemistry, we use energy considerations to figure out what will happen. It turns out that the reaction $\text{Zn} \rightarrow \text{Zn}^{2+} + 2e^-$ is much more energetically favorable than the reaction $\text{Cu} \rightarrow \text{Cu}^{2+} + 2e^-$. Both reactions produce electrons, but the reaction with the zinc is stronger—it's like you have a push on the electrons from the nail and a push on the electrons from the copper, but the push from the nail is stronger. The push from the nail wins out, and in fact causes the reaction at the copper to proceed in the reverse direction. Instead of $\text{Cu} \rightarrow \text{Cu}^{2+} + 2e^-$, what happens is $\text{Cu}^{2+} + 2e^- \rightarrow \text{Cu}$. See Figure 5.8. It's kind of like two downhill slopes competing for the direction a ball will roll. The slope leading from the nail is steeper, so the ball (or in this case, electrons in the outer circuit) rolls toward the copper. And when you push electrons from one place to another, you have a battery.

Figure 5.8



Most batteries work on this same principle. You have two elements (carbon-zinc, nickel-cadmium, and so on) connected by an electrolyte (a substance, usually a liquid, that contains ions that can move freely). There's one point of potential confusion when dealing with chemical batteries. For example, in our vinegar battery, chemists refer to the zinc terminal as being the **cathode**, or positive terminal, and the copper terminal as the **anode**, or negative terminal. These are proper labels when considering what happens to the ions in solution that connect the terminals, but when considering the external circuit (the wires and the bulb), you would properly consider the zinc terminal to be the anode (negative) and the copper terminal to be the cathode (positive) This is because electrons flow in the external circuit *from* the zinc terminal *to* the copper terminal. Figure 5.9 explains it all.

Figure 5.9



If we consider only the external circuit and which way the electrons move, the copper is the positive terminal (cathode) and the zinc is the negative terminal (anode).

If we concentrate only on what's going on in the solution, the copper is the negative terminal (anode) and the zinc is the positive terminal (cathode).

More things to do before you read more science stuff

Time to move on from electrochemistry to how light interacts with matter. I'm going to use acid-base indicators for this purpose, but first I'll start with a couple of simple questions. What makes any object the color it is? Yes, the object reflects certain colors of light, but what is going on with the atoms and molecules of that object to make it reflect only certain colors? How can you change the color of an object? Sure, you can dye a shirt a different color, but what's going on with the atoms and molecules of the shirt to make it change color? Answers in the next section, but try to construct your own answers before reading mine.

SCiLINKS
THE WORLD'S A CLICK AWAY

Topic: Light and Color

Go to: *www.scilinks.org*

Code: MCB014

Next, you will need one or more of the following: phenolphthalein solution (less than an ounce of 1% solution), bromothymol blue solution (about an ounce of 1% solution is enough, and this can be diluted with water in a 1 to 4 ratio and it will work fine), or cabbage juice (just boil a small amount of red cabbage for a few minutes in a cup of water, remove the cabbage, and use the liquid that remains). You will also need sodium hydroxide solution (0.1M to 0.5M)[6], vinegar, and two pipettes (plastic or glass—see Figure 5.10) or eye droppers.

Figure 5.10

CAUTION: Sodium hydroxide in any concentration is a bit caustic and definitely a skin irritant. Wear goggles and gloves.

For what follows, I'll assume you're using phenolphthalein. The procedure works basically the same way for the other two solutions. Pour a small amount of phenolphthalein solution into a clear cup. Add a bit of water. Use one of your pipettes or eye droppers to add sodium hydroxide to the solution, one drop at a time. You should notice a color change to pink that fades to clear. Keep adding sodium hydroxide until the solution remains pink. Once you get a permanent change in color, begin putting vinegar in the solution one drop at a time. Keep doing this until the color of the solution changes back to clear. You can keep alternating sodium hydroxide and vinegar through many color changes.

Using bromothymol blue (BTB) or cabbage juice. If you use these liquids instead of, or in addition to, phenolphthalein, expect to see two or more color changes. Also, with BTB you should add vinegar to start and then use sodium hydroxide to reverse the color changes.

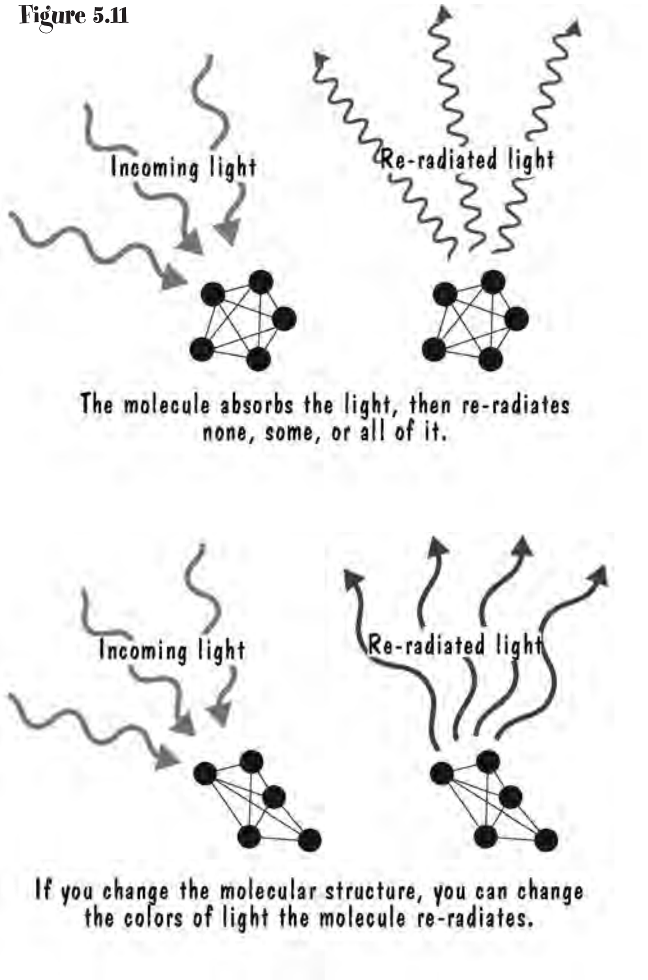
More science stuff

The simple answer to why objects are the color they are is that they reflect certain colors of light and absorb others when white light shines on them. But how do the different colors of light "know" whether they're supposed to be absorbed or reflected? The answer is that reflection of light is more complicated than many people think. Reflected light doesn't just bounce off objects like a ball bouncing off a wall. Reflected light is actually absorbed by the molecules in the object and then radiated back out. The molecules in the object are an integral part of the process, so it should be no surprise that the *structure* of those molecules determines what colors of light are absorbed and what colors are reflected. See Figure 5.11.

---

[6] Recall that 0.5 M refers to the molarity of the solution, so 0.5 M means you have 0.5 moles per liter of solution. You will need this info when ordering the chemical. Often sodium hydroxide comes in tablet form. Follow the product instructions for creating the proper solution, and do be careful. Solid sodium hydroxide might look harmless, but it is caustic.

Figure 5.11



As with chemical reactions in general, the interaction of light with molecules is really the interaction of light with the *electrons* in those molecules. If you change the structure of a molecule, such as by adding or removing certain atoms or changing how the atoms connect, then you can possibly change the colors absorbed and reflected because you are changing where the electrons in the atoms reside.[7]

Now let's turn our attention to acid-base indicators, and phenolphthalein in particular. If you've taken a chemistry course in either high school or college, you have probably used phenolphthalein. It's used to indicate when a solution changes from a pH of below 8.2 to a pH of above 10.0. If the solution in which you have phenolphthalein is acidic or slightly basic, the phenolphthalein is clear. If the solution in which you have phenolphthalein is basic, with a pH greater than 10.0, the phenolphthalein is pink.[8] Ah, but *why* does the phenolphthalein change colors? Must have something to do with change in the molecular structure. Phenolphthalein is actually a weak acid, meaning that it tends to be a proton donor[9] in chemical reactions. When phenolphthalein is in an acidic or slightly

---

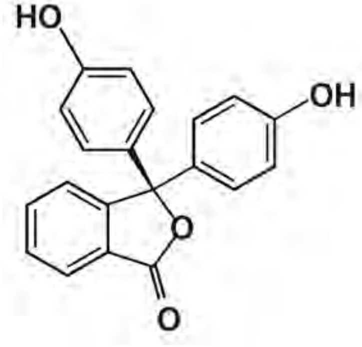
[7] When you change the connections between atoms, you change the distribution of electron energy levels within the atoms, thus changing the "allowed" energies of the electrons. This changes the energy differences between energy levels, which determine the colors of light emitted by an atom. For a review of this process, refer back to Chapter 1.

[8] In really strong, concentrated acids, phenolphthalein is red, and in really strong, concentrated bases, phenolphthalein is again clear. We can safely ignore those situations, because I'm not going to have you mess around with acids and bases that are that strong and concentrated. Not a fan of lawsuits.

[9] Refer to the previous chapter, where I define acids and bases. Acids are proton donors and bases are proton accepters.

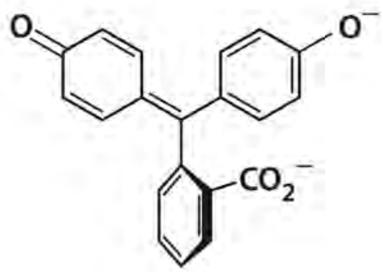
basic solution, the molecule looks like Figure 5.12 and is clear in color.[10] And please don't be intimidated by how complex this molecule is. We're only going to focus on specific parts of the molecule.

Figure 5.12



Notice the two OH combinations that are on one side of the molecule. The bond between the oxygen and hydrogen atoms here is not all that strong, so if some other molecule comes along to entice the H^+ ion to leave this combination, it doesn't take much convincing. Well, sodium hydroxide breaks up into Na^+ and OH^- in solution. The free OH^- ions are just the thing to lure the two H^+ ions away from the phenolphthalein molecule. After the protons leave, and after a bit of reorganization, the phenolphthalein now looks like the drawing in Figure 5.13. This new structure re-radiates the color pink.

Figure 5.13



Let's review the concepts I just explained. Whenever a molecule changes shape or structure, as happens in a chemical reaction, the electron energy levels in the atoms contained in the molecules change. Because light is produced when electrons jump from a higher energy level to a lower energy level, and the energy of the light produced is equal to the difference in energy between the two energy levels, a change in energy levels can possibly change the color emitted in the electron jumps. I used acid-base indicators as an example, but I could have used any chemical reaction in which there is a color change. Chemical reactions involve rearrangement of atoms and new structures, and this often leads to changes in re-radiated light. In fact, a color change is one of the first things students learn to look for as evidence that a chemical reaction has taken place. The reason I used acid-base indicators as an example is that it gives us an opportunity to review acids and bases, and you learned that the indicators themselves are either acids or bases.

SCILINKS THE WORLD'S A CLICK AWAY
Topic: Molecular Shapes
Go to: *www.scilinks.org*
Code: MCB015

Bromothymol blue and cabbage juice work in much the same way as phenolphthalein. As the acidity of the solution they're in changes, their molecular structure changes as well. These different structures, three kinds in each case, re-radiate different colors of light. Of course, there are many more kinds of acid-base indicators, and each one is either a weak, low-concentration acid or a weak, low-concentration base. That's why they interact with the acids and bases

---

[10] Those hexagons in the phenolphthalein molecule might not be familiar to you. They're "rings" of carbon atoms called *benzene rings,* with a carbon atom at each corner. Such rings are common enough in chemistry that you'll see this shorthand notation a lot.

in various solutions. What makes a substance a good acid-base indicator is its ability to change color in a small range of pH values.

One last thing in this section. Why are acid-base indicators *weak* and in *low concentrations*? Wouldn't a strong, concentrated acid or base react even better with the acids and bases in the solution you're dealing with? Yes, it would, and that would be a problem. You want your indicator to indicate through a color change what the pH is of *other* solutions. If your indicator is strong and concentrated, then the pH of the indicator governs the pH of the entire solution.

Even more things to do before you read even more science stuff

Most of the activities I'll have you do in this section represent fairly common experiences, so if you have a good memory for these experiences, you don't absolutely have to get the things I ask. These are cool activities, though, so you might not want to pass up a chance for playtime. First, get a few glow-sticks and a few glow-in-the-dark stickers. Both can be found at your local dollar store. Follow the instructions for using these. How do the instructions differ? Which requires an external source of light? How long does each of these glow? Which is rechargeable?

If you're a baby boomer pack rat, pull out your old "day-glo" posters and your "black" light. If you have no such relic, head to your nearest laser tag facility. This is day-glo heaven. Notice that black lights (which emit primarily ultraviolet light) bring out all sorts of interesting colors, on a poster or your clothes, that weren't there in normal incandescent light or sunshine. What created the new colors and the extra brightness? Was it a change in molecular structure?

Topic: Ultraviolet Light
Go to: *www.scilinks.org*
Code: MCB016

Even more science stuff

Let's start with the last thing I asked you to do (look at things that glow under black light) because that's actually the easiest to understand. Recall that atoms emit light by electrons jumping from a higher energy level to a lower energy level and can absorb light with the result of their electrons jumping from a lower energy level to a higher energy level.[11] This is also true of the electrons in molecules. The only main difference is that the electron energy levels in molecules tend to be a bit more complicated than those in single atoms.[12] The principle is the same.

---

[11] Check out the brief review of this in Chapter 1 of this book or a fuller explanation in Chapter 2 of the first chemistry book in this series.

[12] Keep in mind that when atoms get together to form molecules, that alters what the electrons are doing. Hybrid orbitals form between the separate atoms, and this leads to different energy levels than you would expect from single atoms.

Electrons in molecules can absorb energy, in the form of light or some other source, and move to higher energy levels. Then those electrons can jump down to lower energy levels, with the emission of energy that is sometimes visible light. When you expose certain molecules to ultraviolet light, which is more energetic than visible light, you can excite the electrons in those molecules to energy levels they don't ordinarily attain (with exposure to everyday visible light, that is). Once the electrons reach these high energy levels, you might expect them to just jump back down to lower levels, emitting ultraviolet light, but something a bit different happens. The molecules and the electrons in them lose energy as the molecules go from a situation in which they're vibrating to a situation in which they're not vibrating as much. See Figure 5.14.

Figure 5.14

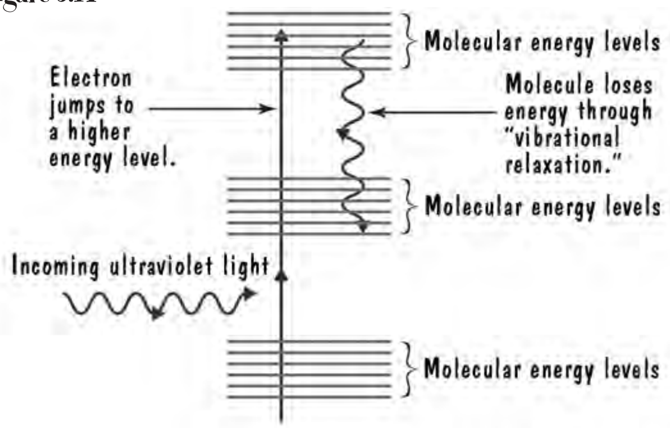
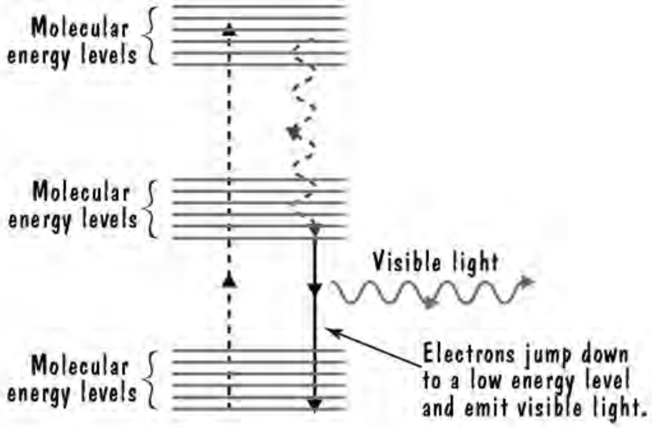


Figure 5.15



The energy jump here is smaller than the energy jump from the input of ultraviolet light shown in Fig. 5.14. This smaller energy transition results in visible rather than ultraviolet light.

Once they've gone through this "vibrational relaxation," the electrons in the molecules now jump down to a lower energy state and emit light. See Figure 5.15.

These new electron transitions often are in the visible light range and also are highly efficient, resulting in bright colors. Because the energy levels available to electrons in molecules act as a fingerprint of the molecules, the colors you see are unique to particular kinds of molecules. To summarize, you shine ultraviolet light on something, causing its electrons to jump to higher energy levels. After the molecules containing those electrons lose vibrational energy, those electrons then jump back down to lower energy levels, emitting visible light. This leads to way-cool day-glo effects and also helps identify elements (see the Applications section, p. 105). There's a special name for this effect—**fluorescence**.

When you stop shining ultraviolet light on a day-glo poster, the poster stops glowing. Not so when you shine visible light on glow-in-the-dark stickers or glow-in-the-dark paint. These things keep emitting light for a long time. Clearly, this is something different from fluorescence. Instead of undergoing "vibrational relaxation" as in fluorescence, the electrons in the molecules undergo a transition to special energy levels in which they're trapped. Transitions from these special energy levels to lower energy levels, with the emission of visible light, are actually "forbidden," meaning they shouldn't happen. But as often happens in the world of quantum mechanics, things that are forbidden actually do happen. I won't go into the math of quantum mechanics that describes all this (As Stuart[13] would say in his demonic voice, "It's ugly!"), but I'll tell you that it has to do with the math telling us that it's impossible for atoms to go from one set of quantum numbers to another set of quantum numbers in a direct transition. Anyway, these so-called forbidden transitions have a low probability of happening. (By now you should realize that probabilities are at the heart of quantum mechanics.) This means that those energy transitions, with the emission of visible light, don't happen very often. The result is that the emission of visible light takes a long time as the electrons stuck in those special energy levels eventually move down to lower energy levels, meaning objects glow for a long time after they've been exposed to the initial energy that excited the electrons in the first place. This process has a special name—**phosphorescence**. Glow-in-the-dark objects undergo phosphorescence.

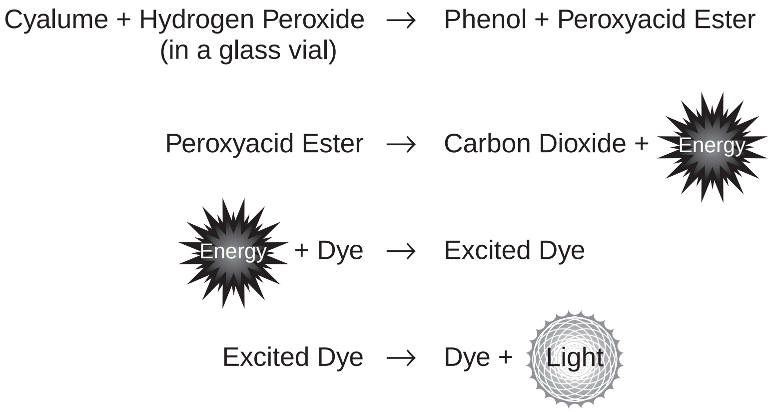
On to glow sticks, which, once you snap them, glow all by themselves for hours. When you snap a glow stick, you break a glass tube and mix different chemicals. Inside the glass tube is hydrogen peroxide, the stuff you use to clean wounds. In the plastic tube, separate from the hydrogen peroxide, is a chemical called **cyalume** along with a fluorescent dye.[14] The cyalume and hydrogen peroxide mix and produce a molecule known as phenol (active ingredient in a popular throat spray) and another known as peroxyacid ester. The peroxyacid ester spontaneously transforms into carbon dioxide while giving off energy. The fluorescent dye then absorbs this energy and gives off the light you see. This process, shown in Figure 5.16 (p. 104), continues as long as there is hydrogen peroxide and cyalume left in the tube. This might seem like a complicated process, but conceptually it's simple. Chemical reactions produce energy, and this energy causes a dye to undergo fluorescence.

---

[13] Someone familiar with *MADtv* can tell you about this obscure reference.

[14] You can cause this dye to glow before you ever snap the stick by shining ultraviolet light on the stick.

Figure 5.16



I know that Figure 5.16 looks a bit complicated. Just think of it this way. The top two chemical reactions result in excess energy. That energy then acts like the ultraviolet light in our day-glo example, exciting the electrons in the dye to higher energy levels (that's the third equation). Then, after undergoing vibrational relaxation, the electrons in the dye jump down to a lower energy level and emit the light you see.

To summarize this section, light interacts with matter by interacting with the electrons in atoms. It's all about electrons being excited to higher energy levels and then dropping down to lower energy levels with the emission of light. This is true even with reflected light, which is really the absorption and re-emission of incoming light. You can alter the light that a substance emits by changing molecular structure (which changes the available energy levels) or by hitting atoms or molecules with energy that puts electrons in energy levels in which they don't normally reside when hit with visible light.

Chapter summary

- Liquids containing ions, referred to as electrolytes, can conduct electric current. Electrochemistry is the study of electric current by ions and of the associated chemical reactions.
- The polar nature of water molecules separates many ionic compounds into ions.
- The flow of ions in an electrolyte can be used to electroplate many metals.
- Chemical reactions between various substances and electrolytes can serve as a battery, causing electrons to flow in an external circuit.
- Two half-reactions take place in batteries, with the stronger reaction driving the direction of flow of electrons in an external circuit.
- Objects and substances absorb some colors of light and reflect others. Changes in the molecular structure of substances often change the colors absorbed and the colors reflected.

- Acid-base indicators are either weak acids or weak bases. Because they interact with stronger acids and bases and change color in a limited pH range, they are useful as pH indicators.
- Fluorescence is the process in which molecules absorb ultraviolet light or absorb energy in some other way and quickly emit visible light.
- Phosphorescence is the process in which molecules absorb visible or ultraviolet light and then emit visible light over a long period of time.

Applications

1. Most applications of electrochemistry involve redox reactions, which we went over in the previous chapter. The reactions in a battery are redox reactions. Another common application is the corrosion of metals, as with the rusting of iron. I′m not going to provide all the equations here that go into the rusting of iron because the process is relatively complicated. I do encourage you, though, to look up rusting on the internet. Although there are quite a few equations, each individual equation is relatively straightforward. There are gains and losses of electrons, plus the role of an electrolyte (in this case, water or saltwater) in helping the transfer of electrons. In this way, rusting is not so different from the operation of the battery you created in this chapter.
2. The following website gives you a chance to see what fluorescence in minerals looks like without having to buy an ultraviolet light. Simply move your cursor over the samples to see what they look like under UV light. *http://mineral.galleries.com/minerals/property/fluoresc.htm*.[15]

   Fluorescence isn′t the only light-related way to identify rocks and minerals. If you heat some minerals, that can cause electrons trapped in high energy levels to jump down to lower energy levels and emit light, in a process known as **thermoluminescence**. As with all light emitted from atoms, the colors are different for different atoms, leading to identification. A third, less reliable method for identifying rocks and minerals is called **triboluminescence**. Some minerals will emit light when crushed, for the same reason they emit light when heated. Of course, you might want to know why we want to identify rocks and minerals in the first place. Well, oil and gas companies like to know what kind of rocks they might be drilling through, as do hydrologists who might be looking for water supplies. Construction engineers are interested in rock and soil types because rock and soil types

---

[15] There′s always a danger that web links expire over time. If this link doesn′t work for you, just put ″fluorescence minerals″ into a search engine and you′ll probably find a site where you can see examples of mineral fluorescence.

affect the design of various projects. And people looking to mine minerals have an obvious interest in where the minerals are.

3. Go into the bathroom, turn off the lights, and look in the mirror as you chew on a wintergreen Life Saver with your mouth open. Triboluminescence in action.
4. We used acid-base indicators to talk about how light interacts with matter, so maybe we should discuss how we use acid-base indicators. One primary use is in a process called titration. Suppose you have one liquid that has a known concentration, or molarity, and another liquid of unknown concentration. If these happen to be an acid and a base, then what we do is add, say, a bit of phenolphthalein to the liquid of known concentration. Then we slowly add the liquid of unknown concentration until the phenolphthalein turns either pink or clear. At that point, we know the pH (or rather the range of pH, because phenolphthalein changes color in a pH range) of the combined liquids and can figure out the concentration of the unknown liquid.
5. Many organisms exhibit bioluminescence, meaning they emit their own light. Contrary to popular opinion, this is not an example of fluorescence or phosphorescence. You already knew that, though, because organisms emit their own light in dark conditions, and fluorescence and phosphorescence require the input of light. Bioluminescence does, however, depend on chemical reactions, and of course the light comes from transitions from high to low energies. Here's how it works in fireflies, which are actually beetles. The reaction depends on adenine triphosphate (ATP), which, as anyone who has studied any biology knows, is involved in carrying energy from one place to another. In a firefly, ATP combines with magnesium ions and a substance called luciferin[16] to form a very high-energy, unstable molecule. This molecule readily combines with oxygen and transforms to a lower energy state, with the emission of light.

---

[16] Besides its other meaning, *lucifer* is Latin for *light-bringer.* Makes sense.

6 Chapter

Half a Life Is Better Than None

Most of chemistry deals with chemical reactions and, hence, with how electrons in atoms behave either in single atoms or when those atoms get together with other atoms. There is a branch, though, that deals with the nuclei of atoms, and that's what this chapter is about.

" ... so the first radioactive isotope says to the second, 'Half my life is over! I just hope the other halves don't go as fast.'"

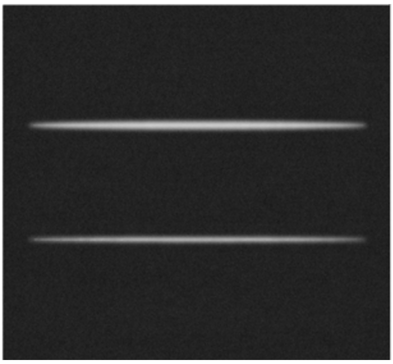
Things to do before you read the science stuff

For most subject areas in basic chemistry, you can base your understanding of concepts on easily observable things such as fizzing and color changes. The scientific concepts that help you understand the observations might be really abstract (ever seen an electron?), but at least you have something concrete (such as balloons exerting a force on a stream of water) to relate them to. Not so with nuclear chemistry. In the next two Things to Do sections, I will mainly have you do a few things that serve as analogies for the concepts I'll introduce later.

Topic: Half-Life
Go to: *www.scilinks.org*
Code: MCB017

With that said, I'll show how consistent I can be by telling you about an activity you can do to observe something directly. Head to your local internet search engine and enter the term **cloud chamber**. You'll find a number of designs for building a cloud chamber. Because the procedure is quite a bit more involved than the kinds of activities I normally ask you to do, I'll leave it up to all of you Bob Vilas out there to build one. For the rest of us, I'll let you know what you can see in a cloud chamber. Even without a radioactive source nearby, you will see occasional "tracks" form in the mist of the chamber, indicating that something very small has passed through the mist. If you put a radioactive source nearby, you'll see many more tracks form. They look something like those in Figure 6.1.

Figure 6.1

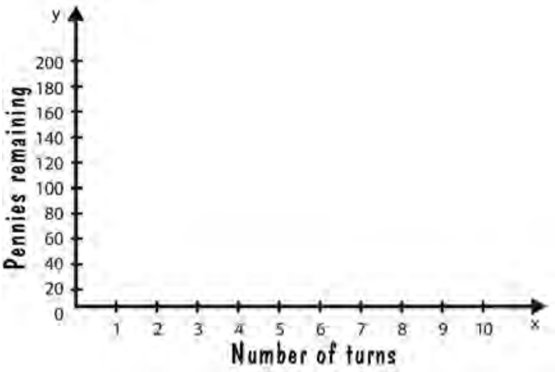


Now on to analogy-type activities. Gather 200 regular M&M'S, 200 pennies, 100 dice, or any other objects that can give one of two or more results when poured out on a table. For example, an M&M can randomly land letter side up or letter side down. A penny can randomly land tails up or heads up. A die can randomly land with between one and six dots facing up. For what you're about to do, I'll assume you're using 200 pennies.

Find a clear, smooth surface. Place all the pennies in a bag and mix them up thoroughly. Then empty the pennies out on the surface. Remove and set aside all the pennies that landed heads up and count the number of pennies remaining on the surface (the ones that landed tails up). Record this number. Place the pennies that landed tails up back in the bag and mix them thoroughly. Pour these out on the surface. Remove and set aside the pennies that landed heads up and count the pennies remaining on the surface (the ones that landed tails up). Record this number. Keep repeating this procedure until you have no more pennies left on the surface. When you're done, make a graph showing the number of pennies left on the vertical axis and the number of turns (each turn is when you throw the pennies on the surface and remove the heads) on the horizontal axis, as in Figure 6.2.

And a quick note if you use dice instead of pennies or M&M'S: Each time you throw the dice on the surface, remove just the dice that are showing two or the dice showing six (or any other number you choose). By removing, on average, only one-sixth of the dice, the procedure will take longer. You'll end up with a smoother graph, though. Also, whatever materials you use, you will get better results if you do the activity four or five times and average the results. Of course, if you have access to a room full of students, you can do the activity ten or more times, with each group doing it once.

Figure 6.2



The science stuff

I'm going to explain what kinds of things happen in radioactivity in this section. In the explanation section after this one, I'll address the all-important question of why radioactivity occurs.

A comment on the *why* question. When reviewing radioactivity for this book, I was reminded that too often in science resources, authors explain what happens without explaining why it happens. If you can only describe occurrences, then you really don't understand what's going on, and you end up only memorizing what happens. If you can figure out a mechanism for the occurrences, though, then you can build a lasting understanding of what's going on. Even though scientists often can only describe what happens when they first encounter a phenomenon, the ultimate goal is a mechanism for the phenomenon and the resultant understanding. You can compare this to mathematics, in which there are rules to follow. Only when you understand the reasoning behind the rules do you understand math.

Before you continue, you might benefit from a quick review of the structure of an atom. There is a central nucleus that consists of protons with a positive charge and neutrons with no charge. Surrounding this nucleus are negatively charged electrons. In this chapter, although electrons play a part as in the rest of chemistry, we'll find that they play a relatively small part in nuclear reactions. We'll also discover that there is a structure to the nucleus that goes beyond just the presence of protons and neutrons.

A cloud chamber contains supercooled, supersaturated alcohol vapor. If you send any tiny, charged object into this vapor, that charged object will cause ions to form in the vapor. These ions serve as places on which the vapor can condense (turn to liquid) and form tiny little "clouds." If the charged object moves really fast, it will leave trails of tiny clouds, which form tracks.

The charged objects that cause tracks in a cloud chamber are actually particles emitted from the nuclei of atoms. The nuclei of some atoms undergo changes, such as a neutron turning into a proton plus an electron, a proton turning into a neutron with the disappearance of an electron, a nucleus spitting out a neutron, and a nucleus spitting out an alpha particle (the nucleus of a helium atom—two protons and two neutrons). There is a whole range of particles you might or might not have heard of in these nuclear changes, such as neutrinos, anti-neutrinos, gamma particles, and positrons. As I said earlier, I'll explain why in the world such things might happen in a later section. For now, I'll just focus on one such change to illustrate the kinds of symbols and terminology used.

In alpha particle emission, a nucleus simply emits two protons and two neutrons (bound together). An example is the element radon-222, which has an atomic number of 86 (86 protons) and an atomic mass of 222 (222 total protons and neutrons together).[1] For this and other radioactive transitions, you can write an equation similar to a balanced chemical equation. Here it is:

$$\text{radon} \rightarrow \text{polonium} + \text{alpha particle (helium nucleus)}$$

$$^{222}_{86}\text{R} \rightarrow {}^{218}_{84}\text{P} + {}^{4}_{2}\text{He}^{2+}$$

or

$$^{222}_{86}\text{Rn} \rightarrow {}^{218}_{84}\text{Po} + \alpha$$

Notice that I've written each atom or particle with a new notation. The subscript in front of each symbol gives the atomic number (the number of protons in the nucleus) and the superscript in front of each symbol gives the atomic mass (the total number of protons and neutrons in the nucleus). This is common notation in radioactive decay, so you can keep track of the number of protons and neutrons. Also notice that because the radon lost a couple of protons, it transformed into a new atom—polonium.[2] The symbol for the helium nucleus is written ${}^{4}_{2}\text{He}^{2+}$ with 2+ as a superscript because it has a charge of +2. Remember

---

[1] Just for practice, you might want to take a look at the Periodic Table and confirm these numbers.

[2] The number of protons in an atom determines its identity. You can change the number of neutrons or electrons and still have the same element, but changing protons results in a different element.

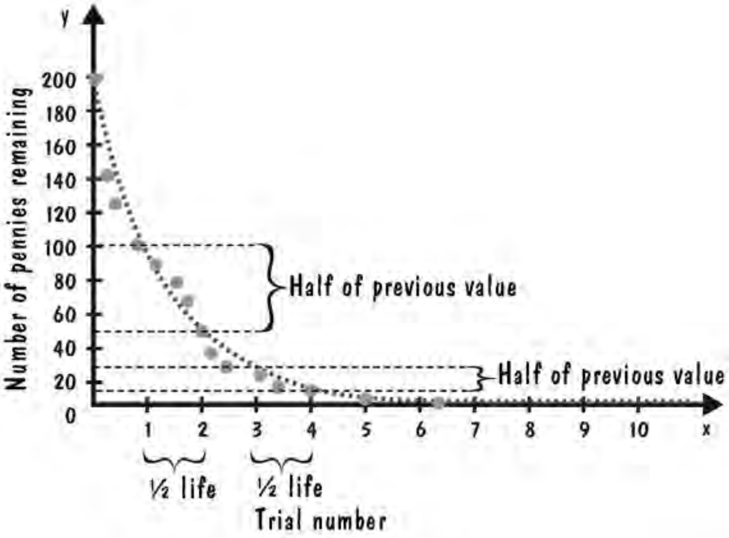
that it's just the nucleus, with the electrons that would make it a helium atom removed. Finally, I've replaced the symbol for the helium nucleus with the Greek letter α (pronounced *alpha*) because an alpha particle *is* a helium nucleus. The process above is called **radioactive decay** because the original radon lost part of its nucleus (the α particle) and became polonium, which has fewer overall particles in the nucleus. And yes, an α particle would leave a track in a cloud chamber.

SCILINKS
THE WORLD'S A CLICK AWAY

Topic: Radioactive Decay Process
Go to: *www.scilinks.org*
Code: MCB018

Next let's look at the activity I had you do with pennies, M&M'S, or dice. The graph you made should have looked something like Figure 6.3.

Figure 6.3



This graph is typical of any kind of situation in which a certain percentage of the number of existing objects decreases or increases with time. The name for this is **exponential decay** or **exponential growth**. For an example of exponential growth, look at any population growth curve and notice that it's the same curve as in Figure 6.3 with a different orientation. An example of exponential decay would be the amount of aspirin, or any other drug, left in the body over a period of time.

One characteristic of exponential decay is that the population of objects halves itself in a regular time interval. In the case of pennies or M&M'S, the number of remaining objects cuts in half about every turn. If you use dice and only remove, say, the dice that reads "3," then the population of dice halves itself over a time of several turns. Because all exponential decay curves do this, we use the term **half-life** to describe how fast the decay happens. Notice that two different half-lives are shown on the graph. In the second, the drop-off in number of pennies remaining is less than in the first. That's because in one half-life approximately half of the *existing* pennies are removed. Of course, the time for one half-life (one trial in this case, several trials if you use dice) is the same value throughout.

Because I seldom have you do activities just for the fun of it, you probably have guessed that the simulation models radioactive decay, and it does. When one atom changes into another via emission of a radioactive particle and a change inside the nucleus, the decay is exponential. The number of atoms that haven't decayed (like the number of pennies remaining) exhibits a graph similar to Figure 6.3, and there is an associated half-life. In fact, the half-life of a radioactive element is useful for determining how long the element will remain significantly radioactive—useful in everything from medicine to nuclear reactors.

There's a mathematical relationship, which I'll outline for you, that describes exponential decay. You start with a relationship between the change in a quantity (number of pennies removed, amount of a radioactive element that decays) and the amount of the quantity present at any time. In words, it looks like this:

$$\text{change in quantity} = (\text{some constant})(\text{amount of material remaining})$$

So, the amount of change, or the amount of stuff you lose, is directly related to the amount you have at any time. In symbols, we can write it as

$$\text{change in } N = (\text{a constant})\cdot N$$

where N represents the number of things you have at any time. This makes sense because how many things we lose in a given time period depends directly on the number of things we currently have. Now we invoke the magic of calculus (no, I'm not going to do any, so don't put the book down) to come up with a mathematical relationship that tells us how many things we have (this is N) after a time period (t), given that we start with a certain number of them (we use N_0 to represent that number we start with). Here's that relationship:

$$N = N_0 e^{-\lambda t}$$

In that equation as applied to radioactive decay, N stands for the number of atoms that have not yet decayed after a time (t), N_0 stands for the number of atoms you started with, e stands for a special number associated with natural logarithms that is approximately equal to 2.718, and λ stands for a decay constant that is different for different atoms. From this, you can determine the half-life of any radioactive element. It's a fairly simple derivation, but since we're not going to do any calculations with the above relationship, I'll spare you. Just remember that we can use the above equation to figure out how much of a substance has not decayed after a certain time, and that's a good thing to know. For example, we need to know when a radioactive substance used in a nuclear reactor is down to insignificant radiation levels for the purposes of storage. If you plan to use a radioactive dye (see the Applications section, p. 120) in the human body, you need to know that it will last long enough for you to make the necessary observations of body parts. All radioactive elements have distinctive half-lives, and it's pretty simple just to look them up when needed. Nice to know you don't have to make your own measurements each time.

Two final things before moving on to why some elements are radioactive. Because there are billions of individual atoms in a sample of radioactive material,[3] you never, in a practical sense, run out of radioactive material. You just keep taking halves of halves of halves and never getting to zero in any amount of time that matters to everyday life. So, if your understanding of half-lives is that after two half-lives everything is gone, then you don't have the concept.

That brings up the second thing, which is that there are acceptable levels of radiation. We are constantly bombarded with naturally occurring radiation from outer space and natural elements in the Earth. You've been bombarded with particles from radioactive decay since you were born. Of course, even naturally occurring radiation can be harmful. In my part of the country, it's wise to check the levels of radon underneath your home because its radioactive particles can cause health problems.

Just a few more things to do before you read more science stuff

Some of the activities in this section might seem to have nothing to do with radioactivity, but they do. It's important to try these activities so you can understand why elements are radioactive and how we use radioactivity to generate power. The first thing to do is pull two magnets apart and then let them snap back together, as in Figure 6.4 (p. 114). Then think about any energy transformations

---

[3] Just an aside if you're interested. There are exact numbers for the half-lives of elements, and we can get those exact numbers precisely because there are so many atoms in a radioactive sample. When considering individual atoms, though, we cannot predict exactly when the atom will decay or exactly how much energy the escaping particles will have. This is because of the uncertainty principle, discussed in Chapter 3. We cannot simultaneously know the energy of an interaction and the time over which the interaction takes place, so there is an inherent uncertainty in when this happens. As I said, though, the incredibly large number of atoms smoothes out this uncertainty and makes it possible to calculate accurate determinations of half-lives.

Figure 6.4

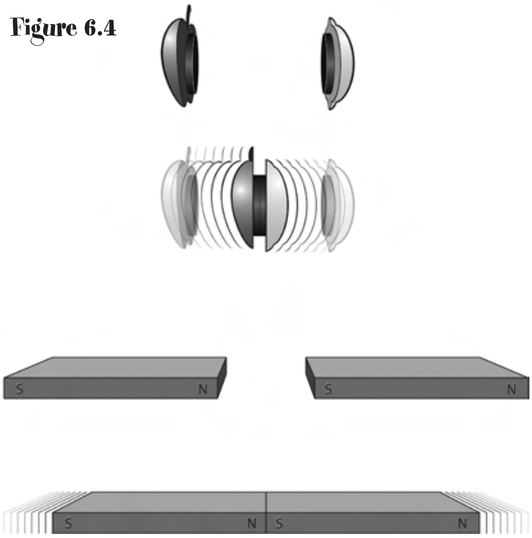
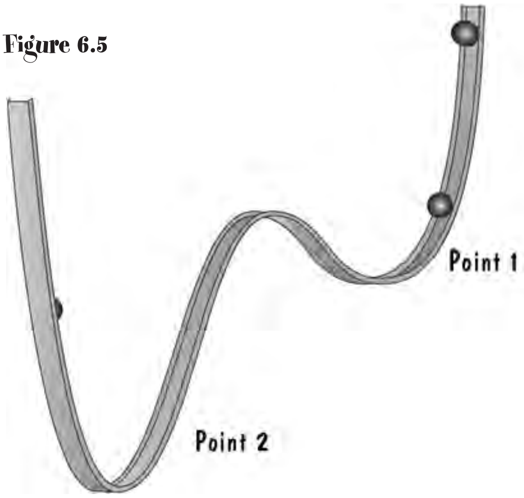


Figure 6.5



that might have taken place. When did the magnets have more energy—when they were apart or together?

Your next task is to get a marble or other ball and a length of Hot Wheels track or some other track on which the marble can roll freely. Shape the track as in Figure 6.5.

Start the marble at point 1. Push it up over the hump and let it come to rest at point 2. Where do you think the marble had more energy, at point 1 or at point 2? Did the marble ever have more energy than it did at these two points?

As you're wondering what in the world these activities have to do with nuclear chemistry, ponder this question: What holds the nucleus of an atom together? Shouldn't all those positively charged protons push one another apart? Next, take a look at the Periodic Table and see what happens to the number of protons and neutrons in an atom as the atoms increase in atomic number. In the smaller atoms, the numbers are equal. What happens after they are no longer equal?

Finally, I want you to watch a video. The video shows a bunch of Ping-Pong balls set on a bunch of mousetraps. One ball thrown onto the bunch causes all sorts of commotion. You can find such videos many places on the web, but the best one I found is at *http://natureofthechemicalblog.blogspot.com/2007/10/atomic-mousetraps.html*. If that link disappears for some reason, just plug "mousetraps chain reaction" into an internet search engine and you'll find a similar video.

More science stuff

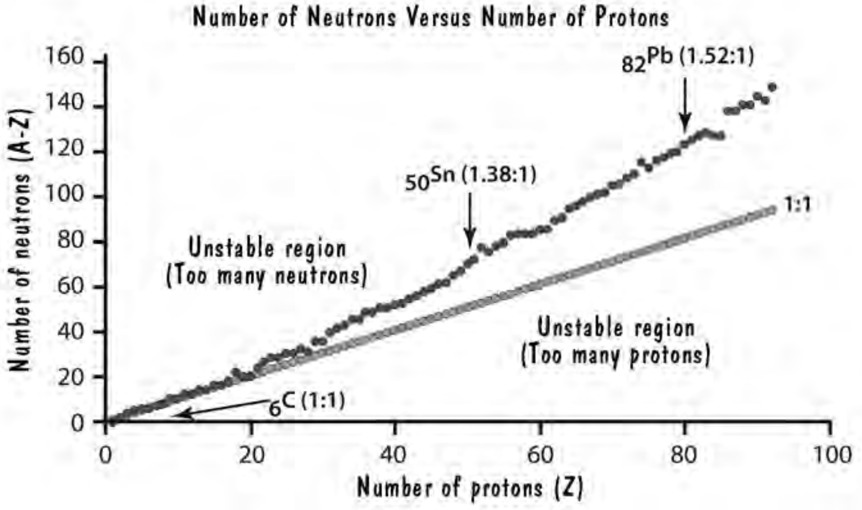
I encourage you to take your time in this section. These are not necessarily easy concepts or concepts you will encounter unless you become a physics or chemistry major in college. However, I think they will greatly increase your understanding and perspective and greatly improve your ability as a teacher.

One of my pet peeves is that we hide from people the kinds of things in this chapter because the ideas aren't traditionally taught at lower levels. Yet these things are understandable and really help create a worthwhile view of science. No complicated math, so please don't shy away from this stuff.

Let's address the questions in the middle of the previous section first. It does seem that the positively charged protons in a nucleus should repel one another and break apart the nucleus. Remember that neutrons have no charge, so they don't exert an electric force. There's another force in play, though. Nucleons (protons and neutrons) exert a very strong force on one another when they get close. It's called the strong force (makes sense) or the **strong nuclear force**. This force is stronger than electric forces or gravitational forces but doesn't have an effect unless the nucleons are almost on top of one another. So, while the strong force is limited to the realm of the nucleus, it is more than enough to overcome the repulsive electric force between protons, and the nucleus holds together.

If you take a look at the Periodic Table, you'll notice that up to the element sulfur, with the exception of hydrogen and isotopes, the numbers of protons and neutrons in the element are equal (the atomic mass is roughly twice the atomic number). After that, the elements begin to have more neutrons than protons. Why? Because when you add neutrons to a nucleus, you increase the number of particles exerting a strong nuclear force without increasing the positive charge that tends to push the nucleus apart. So, the more neutrons, the better, right? Evidently not, because most of the heavier elements have protons and neutrons in the ratio of about 2 to 3 instead of having way more neutrons than protons. Figure 6.6 is a graph that you'll find in lots of chemistry resources; it shows how the ratio of protons to neutrons deviates from a ratio of 1 to 1.

Figure 6.6

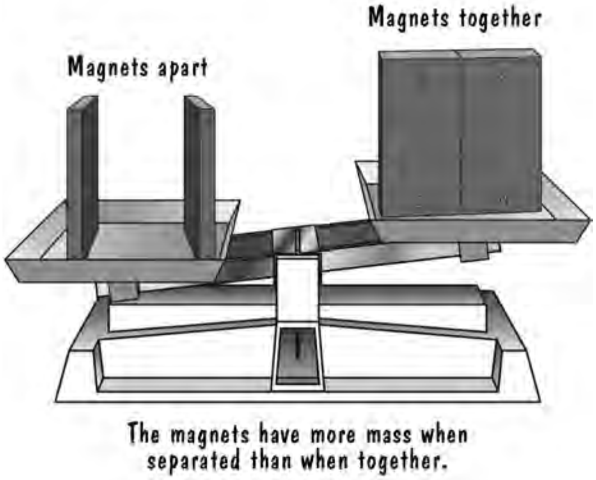


Okay, so what's going on? Why don't atoms have lots and lots of neutrons to create a strong bond and stability? The answer lies in the fact that neutrons and protons have a lot in common with electrons. No two electrons can have the exact same set of quantum numbers, and this leads to the whole concept of "shells" in which the electrons reside. Tap that memory again and recall that only two electrons could fit in the lowest energy shell, two in the next highest energy shell, six in the next highest energy shell, and so on. If you have an atom with shells completely filled with electrons, the next electron you try to add to this atom has to reside in the next highest energy level. That electron is typically easy to remove from the atom.

It turns out that neutrons and protons also have energy shells in which they reside. Therefore, the stability you get by adding extra neutrons and increasing the strong nuclear force can be offset by the fact that you can end up with an arrangement in which you're adding neutrons to higher, unfilled energy shells. Of course, as the energy of the shells gets higher and higher, adding more neutrons will definitely not be stable because all systems tend toward the lowest energy possible. So, we end up with an "optimal" ratio of protons to neutrons. Nuclear configurations that stray too far from that optimum tend toward nuclear change and radioactivity.

Now I'm going to tell you about a strange concept that's necessary for understanding radioactivity and other nuclear reactions. That concept is the equivalence of mass and energy. Mass can transform into energy, and vice versa. This is part of Einstein's theory of special relativity and is the source of that famous equation $E = mc^2$. Let's apply the theory to the activities you did in the previous section. When the two magnets are apart, we say that they have potential energy due to their separation. When you release the magnets, this potential energy transforms into kinetic energy (the motion of the magnets) and, finally, into heat and sound energy when the magnets collide. Just looking at the beginning and end of this process, you'll see that the magnets lose energy. They begin at rest with a certain amount of potential energy and end at rest without that energy. Because energy and mass are equivalent, then the lost energy of the magnets should show up as a loss of mass of the magnets. In fact, they do lose mass, as illustrated in Figure 6.7.

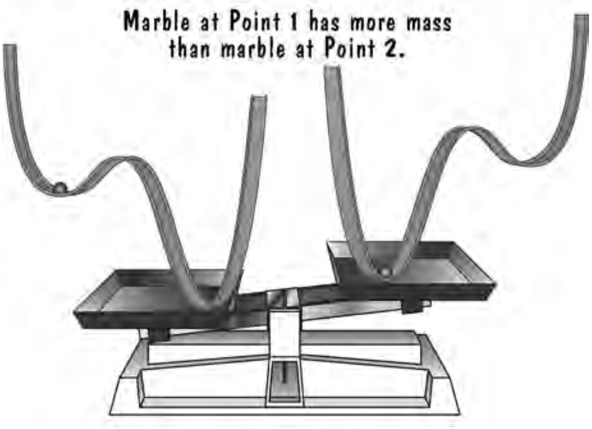
Figure 6.7



This figure is misleading and is not to be taken literally! No matter how sensitive a scale you have, you will not notice a difference in mass between magnets that are apart and magnets that are together. The difference in mass is way too small to measure in a conventional way.

The situation is similar with a marble on a track. When the marble is at Point 2, it has less energy than it does at Point 1.[4] Therefore, the mass of the marble at Point 2 is less than its mass at Point 1. Again, this mass difference isn't something we can actually measure with a scale, but picturing it that way helps (Figure 6.8).

Figure 6.8



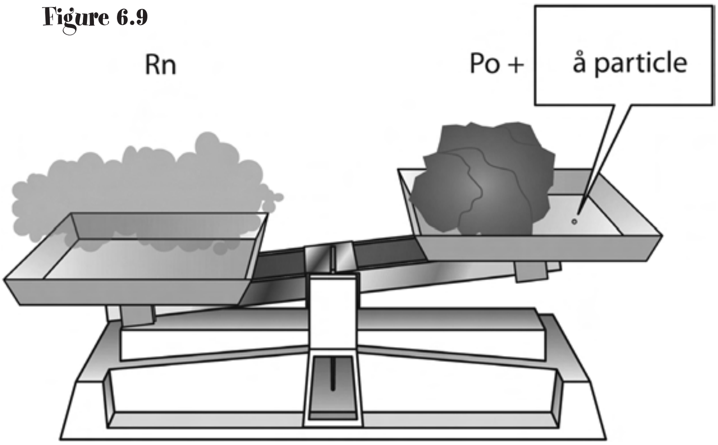
Moving on to radioactive decay, take a look at our example of alpha particle emission.

radon → polonium + alpha particle (helium nucleus)

$$^{222}_{86}\text{Rn} \longrightarrow ^{218}_{84}\text{Po} + ^{4}_{2}\text{He}^{2+}$$

Radon-222 is unstable and spontaneously emits an alpha particle. The resulting polonium is more stable than the radon and has a lower energy state. In fact, there is a transformation of mass into energy (the kinetic, or motion, energy of the alpha particle accounts for most of this), so the radon has a greater mass than the sum of the masses of the polonium and the alpha particle. Cue the scales again in Figure 6.9.

Figure 6.9

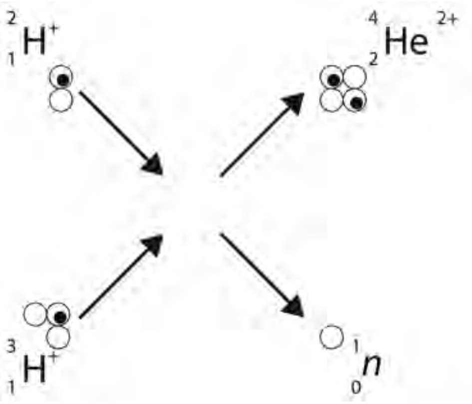


On to another nuclear interaction. Let's take a nucleus of deuterium (hydrogen with a neutron) and a nucleus of tritium (hydrogen with two neutrons) and slam

---

[4] This energy, which is gravitational potential energy, is actually shared with the Earth. Without the Earth, there is no difference in gravitational potential energy between Points 1 and 2.

Figure 6.10

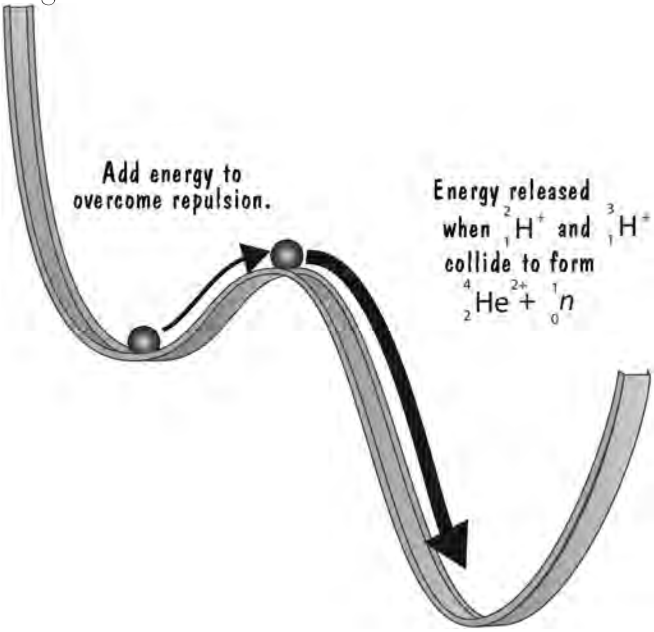


them together. As they approach each other, we have to overcome the electric repulsion caused by both of them being positive. Once they're close enough, the strong nuclear force takes over and they bind together to form a helium nucleus. In the process, a neutron is ejected. Look at Figure 6.10.

Because we have to overcome the electric repulsion, this collision is a lot like pushing a marble up a bump in a track in order for it to slide down into a lower depression, as in Figure 6.11.

The bottom line in all of this is that the mass of the deuterium and tritium is greater than the mass of the final helium nucleus and the extra neutron—this mass is converted into energy. Even though only about 0.4 percent of the mass is converted into energy, that's a lot of energy.[5] This process is known as **fusion**, and it's the nuclear reaction that fuels our Sun and other stars. In stars, this fusion is known as **hydrogen burning** (yes, fusion is what powers our Sun). In warfare, this fusion is known as the hydrogen bomb. In those cases, there are a couple of intermediate steps that I left out, but the principle is the same.

Figure 6.11



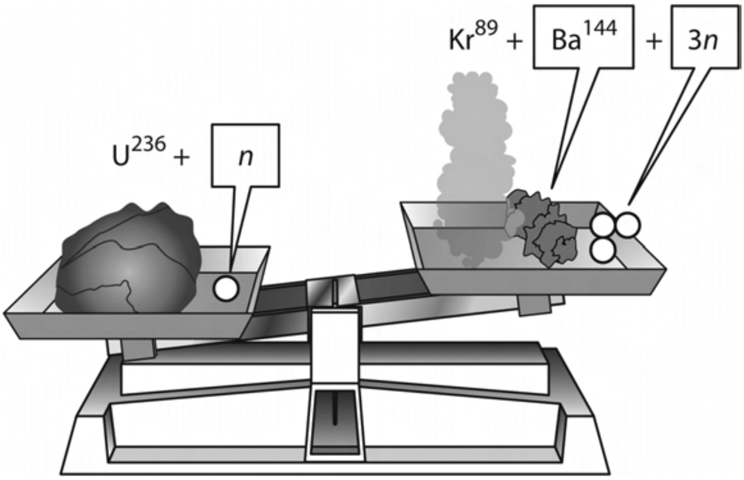
Recall that sometimes adding neutrons and bringing protons and neutrons together results in more stability (as with our fusion example), and sometimes having too many neutrons and protons can result in instability. Uranium-235 is relatively stable, but if you fire a neutron at it, it captures the neutron and transforms into uranium-236. This new isotope isn't

---

[5] For example, if you have a tenth of a gram total of deuterium and tritium, and 0.4 percent of that mass is converted to energy, you get 9×10^{12} joules of energy (I used $E = mc^2$ to figure that out). That's easily enough energy to supply the typical household for about three years.

stable and spontaneously breaks apart into two smaller atoms, one possibility being krypton-89 and barium-144. Three neutrons also break off. If you were to determine the mass of all these things, you'd find that the mass of the original U-236 plus the mass of the incoming neutron is greater than the mass of the krypton plus the mass of the barium plus the mass of the three released neutrons (Figure 6.12).

Figure 6.12



Again, mass has transformed into energy. When nuclei break apart rather than come together, it's called **fission**.

Here's where the Ping-Pong ball video comes in. Each mousetrap that you trigger releases its Ping-Pong ball along with a flying mousetrap. Each of these then triggers another mousetrap. With each release causing the release of at least two more traps, the traps begin to spring at a fast rate. This is known as a **chain reaction**. In our fission example, each of the three released neutrons can plow into a fresh U-235 atom and create another fission reaction. As with the mousetraps, the number of fission reactions increases rapidly (exponentially, in fact). If you have enough concentrated uranium, the chain reaction can result in a bomb. Hence the concern when certain countries start producing enriched uranium. If you can limit the number of neutrons that get produced, though, you have a controlled reaction that is suitable for a nuclear reactor. I'll discuss that in the Applications section (p. 120). I should also add one more thing. The products of an initial fission are often unstable and thus undergo radioactive decay. This leads to all that residual radiation that is potentially so dangerous in both nuclear weapons and nuclear power plants.

Topic: Fission/Fusion
Go to: *www.scilinks.org*
Code: MCB019

With most things in chemistry, if you understand what happens with energy, you understand a lot. If you understand the energy involved in radioactivity and nuclear reactions, you understand what's going on and why. Add in the conversion of mass to energy, and the picture is more or less complete. So, maybe the concepts in this section aren't so difficult after all.

Chapter summary

- Unstable nuclei decay spontaneously into more stable nuclei, usually with the release of various particles. This is known as radioactive decay.
- Radioactive elements have a half-life, which is the time it takes for half of a radioactive substance to decay. The substance doesn't completely decay after two half-lives because subsequent half-lives tell you how long it takes for half of the remaining material to decay.
- The strong nuclear force overcomes electric repulsion and holds neutrons and protons together in a nucleus.
- Protons and neutrons reside in energy shells, just as electrons do. This fact limits the number of protons and neutrons that a nucleus can have and still remain stable.
- There is an equivalence between mass and energy. Whenever a system transfers energy elsewhere, that system loses mass. Energy gains result in an increase in mass.
- Fusion is the joining of small nuclei to form larger nuclei along with the release of energy and/or particles. The energy released in the reaction is much greater than the energy needed to cause the reaction.
- Fission is the splitting of an unstable nucleus into smaller, usually more stable nuclei. Energy and/or collisions with particles cause fission, which results in the release of more energy and/or particles.

Applications

1. People sometimes think of radioactive material as something created by mad and crazy scientists. Although scientists can in fact produce new radioactive elements in the laboratory, the vast majority of radioactive materials we use occur naturally in the Earth. When you hear about gathering, say, fissionable uranium or plutonium, that doesn't happen in the laboratory. We mine these elements just as we mine coal, gold, and silver.
2. You've no doubt been to, or had a friend go to, a radiology department at your local hospital. Radiology used to consist of taking X-ray photographs but now includes much more. For example, some radioactive elements have an affinity for certain human organs or body parts. Radioactive iodine likes to hang out

in bones, so you can inject the stuff into a person, wait for it to accumulate in bone tissue, and then detect the radiation in order to find things such as stress fractures. Of course, you want radioactive elements with short half-lives for this process. Too many beta particles are not good for you!

3. Nuclear reactors are simply controlled fission. The reactor core has a large number of graphite rods that absorb neutrons and keep a chain reaction from happening. The rods can be added or removed to regulate the heat output of the reactor, and the heat output is what the reactor uses to create electricity. The core heats water, which turns to steam. The steam spins a turbine to generate electricity. So, a nuclear reactor is just a fancy way to heat water. Fusion reactions produce a lot more energy than fission reactions, but so far engineers haven't found a way to control a fusion reaction. Once that happens, it's cheap energy for everyone.

4. When carbon-based organisms die, they have roughly the same amount of radioactive carbon-14 in their systems as is present in the atmosphere. Over time, this carbon-14 decays into carbon-12. By measuring the amount of carbon-14 in a long-dead organism and knowing the half-life of carbon-14, we can determine how much has decayed and, therefore, how long ago the organism died. Not surprisingly, this is known as **carbon dating**, which has been a powerful way to date fossils for many years.

Topic: Carbon Dating
Go to: *www.scilinks.org*
Code: MCB020

5. In this chapter, I focused on alpha particle decay, fission, and fusion. There are, however, many other kinds of nuclear interactions that count as radioactive decay. Listed below are a few more.

    - Beta particle emission—A neutron turns into a proton and an electron. The new proton stays in the nucleus and the electron is emitted from the atom, and the electron is called a beta particle. Another subatomic particle, known as an antineutrino, is also produced.[6] An example is the decay of carbon-14 into nitrogen.

      carbon-14 → nitrogen + electron (beta particle) + antineutrino

      The electron, or beta particle, compensates for the increase in positive charge in the nucleus by the transformation of a neutron into a pro-

---

[6] For many years, scientists did not know of the existence of neutrinos and antineutrinos. Because they couldn't keep track of all the energy in this reaction and they believed strongly in the principle of conservation of energy, scientists invented the neutrino as a virtually undetectable particle that carried away some of the energy. All experiments since then have confirmed the existence of such particles.

ton. We don't expect positive charges to appear out of nowhere. The atomic mass remains at 14 because, although a neutron turns into a proton, the total number of protons and neutrons remains the same.

- Positron emission—A proton turns into a neutron plus a positively charged electron, known as a positron or beta-plus particle. As with electron emission, there's another particle included, this time a neutrino instead of an antineutrino. An isotope of fluorine decays into oxygen as follows:

fluorine-18 → oxygen + positron + neutrino

- Gamma particle emission—An atom emits a high-energy photon (particle of light) called a gamma particle (γ). No transformations take place in the nucleus except that it loses the energy the γ takes away.

excited nickel atom → lower-energy nickel atom + high-energy photon

This isn't a complete list of radioactive decay (electron capture by a nucleus and neutron emission are not included), but it's enough for you to get the idea of the transformations involved.

7 Chapter

A Little Organic

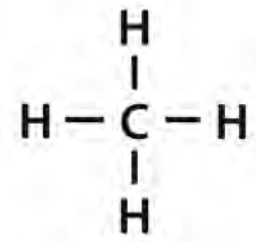
You already know that the study of carbon atoms and the molecules they make has its own special place in chemistry. The reason for this is that carbon atoms combine to make a large list and wide variety of molecules. The reason for the wide variety is the sp^3 hybridization of the four valence electrons in carbon (see Chapter 3).

It would be impossible to investigate even a fraction of organic chemistry in this book, so I'll just go over a few things that will complement the limited organic chemistry in the first book. In other words, don't get upset because this isn't a comprehensive chapter on organic chemistry. In fact, you might not find this chapter all that useful (a few reviewers didn't!). Just look at it as a chance to find out a bit more about all things organic, and you can certainly use this chapter as an application of how hybrid orbitals behave. As with various other material in this book, I am not assuming you will necessarily have an opportunity to teach these concepts to students. You never know, however, when one of your students surprises you with a question about this stuff. It would be nice to know what the heck the student is talking about. Besides, Brian already did the artwork for this chapter, and he'd be really mad if I didn't use it.

Things to do before you read the science stuff

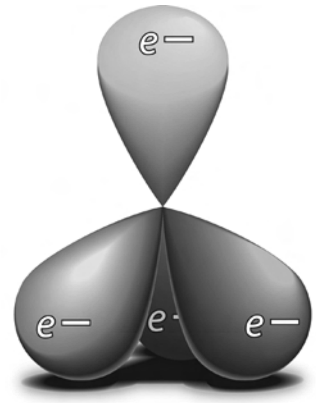
If you thought you could just forget all that stuff on bonding in Chapter 3, think again. I'm going to have you try to apply those concepts to a few organic molecules. Don't worry, I'll provide hints. The first molecule to consider is methane. Its formula is CH_4, with one carbon in the center attached to four hydrogen atoms. We can draw the structure as in Figure 7.1. In that figure, you can see that carbon forms a single bond with each hydrogen.

Figure 7.1



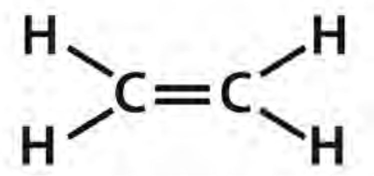
Your job is to figure out what the orbitals look like in these bonds. In other words, try to come up with a sketch similar to Figure 3.32 (p. 65). No, you don't have to worry about distortion of the molecule, as in Figure 3.32, but try to figure out what the hybrid bonding orbitals between the carbon and the hydrogens look like. Your hint is that the carbon atom exhibits sp^3 hybridization, so its orbitals look like Figure 3.24 (p. 59), which is reproduced here as Figure 7.2.

Figure 7.2



So your first task is pretty easy. Just determine what it looks like when four hydrogen atoms, with *s* orbitals, bond to the four sp^3 carbon orbitals in sigma (straight-on) bonds.

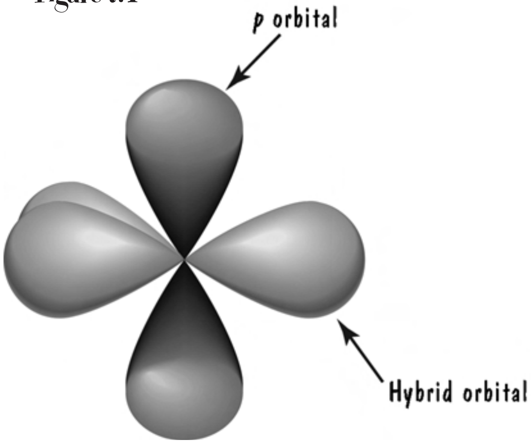
Figure 7.3



The next task is a bit more difficult. Look at the structure of ethene (C_2H_4) shown in Figure 7.3.

Notice that there is a double bond between the two carbon atoms, and recall from Chapter 3 that in a double bond you have one sigma bond and one pi bond (a sideways bond between *p* orbitals). It will also help you to know that in ethene the carbon molecules have sp^2 hybridization. That means that each carbon atom has three hybrid orbitals with one *p* orbital left over. This hybridization looks like Figure 7.4.

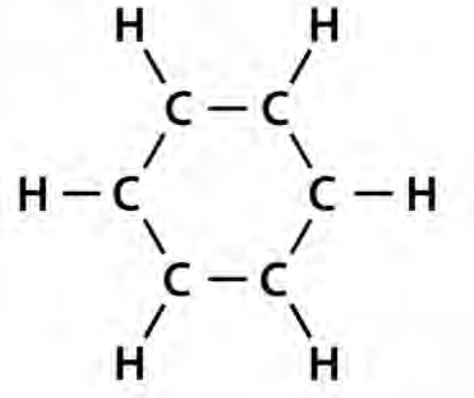
Figure 7.4



After you've figured that molecule out, or at least given it a shot, here's an extremely difficult one. In the molecule benzene (C_6H_6), the six carbon atoms bond together in a hexagon, and the six hydrogen atoms each bond to one carbon atom. The structure looks like Figure 7.5.

Now for a hint or two. The carbon atoms have sp^2 hybridization, just as in ethene. That means three hybrid orbitals and one regular *p* orbital. The carbons each make a single bond with a hydrogen and make single bonds with each carbon on either side. If you get that figured out, then ask yourself what might happen with the regular *p* orbitals that are left over. And no, I really don't expect complete success unless you already know the answer!

Figure 7.5



The science stuff

That was fun, wasn't it? Sort of like you're back in high school or college with a homework assignment that's really difficult. The nice thing about this homework assignment is that I'm going to provide the answers. The orbital diagram for methane (CH_4) is shown in Figure 7.6.

Figure 7.6

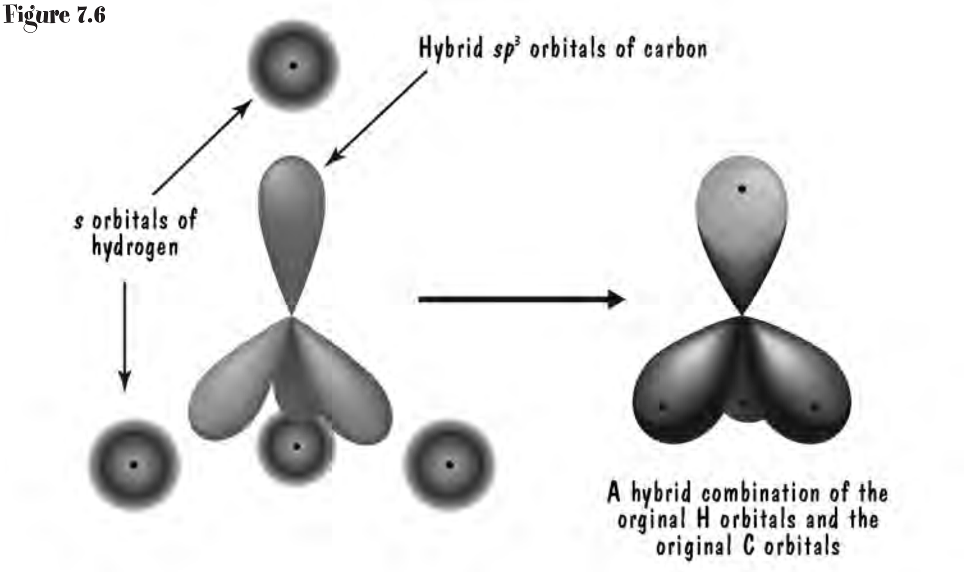
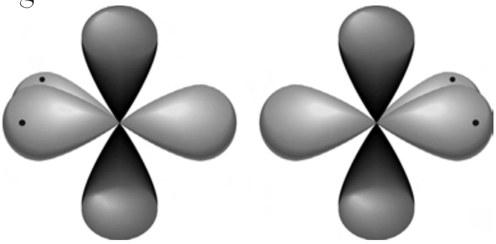


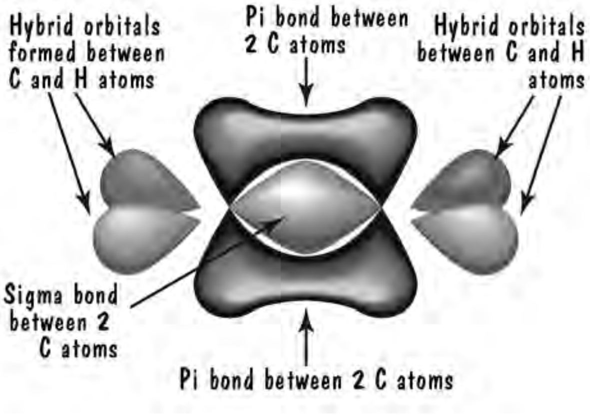
Figure 7.7



Now for ethene (C_2H_4). Figure 7.7 shows the situation before the two carbon atoms bond with each other. There are already bonds with carbon and hydrogen that are the same as the bonds shown in Figure 7.6.

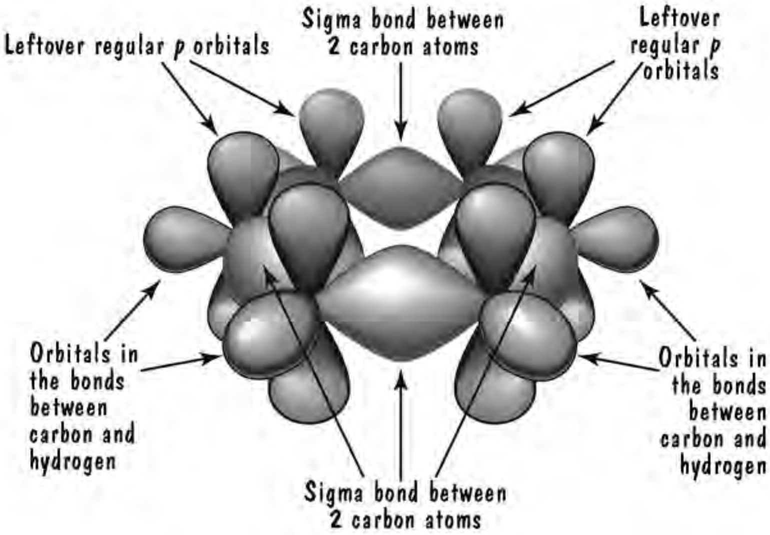
Now we bring those two molecules together and form a double bond between the carbon atoms. The remaining sp^2 hybrid orbitals form a sigma bond, and then the regular *p* orbitals form a pi bond (a sideways bond). The result is Figure 7.8.

Figure 7.8



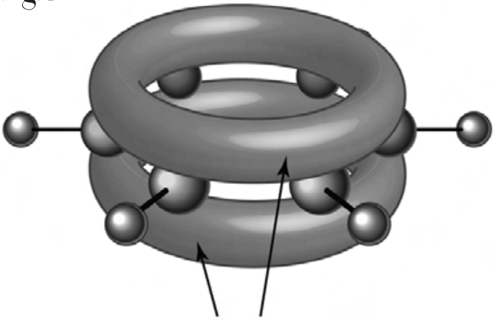
So, we have the double bond between the C atoms that consists of one sigma bond and one pi bond. Now for the final task. I told you benzene forms a ring. The bonds between the carbon atoms are single bonds, as are the bonds between the carbon and hydrogen atoms. It might not be too difficult for you to determine that with those bonds formed, you would get something like Figure 7.9. Notice that we have those leftover *p* orbitals that have one electron each. Shouldn't they just form pi bonds, as in ethylene? You would think so, but consider how that might happen. Two adjacent *p* orbitals would form a pi bond, meaning that with six carbon

Figure 7.9



atoms, you would have three pi bonds between adjacent atoms. Certainly possible. But here we invoke the symmetry of the situation. All the carbon atoms in this ring of carbons are equivalent, with nothing special about any single one or any pair. So one has to ask,[1] "Why would a carbon form a pi bond with the carbon on one side of it when the carbon on the other side looks exactly the same?" The answer is that it doesn't work that way. All of the *p* orbitals in the **benzene ring** combine into one huge hybrid pi orbital, in which one electron from each carbon in the ring resides. This total of six electrons is shared by all of the carbons in the ring. We can represent that as shown in Figure 7.10. This kind of orbital is known as a **delocalized orbital** because it's not localized around one or two atoms.

Figure 7.10

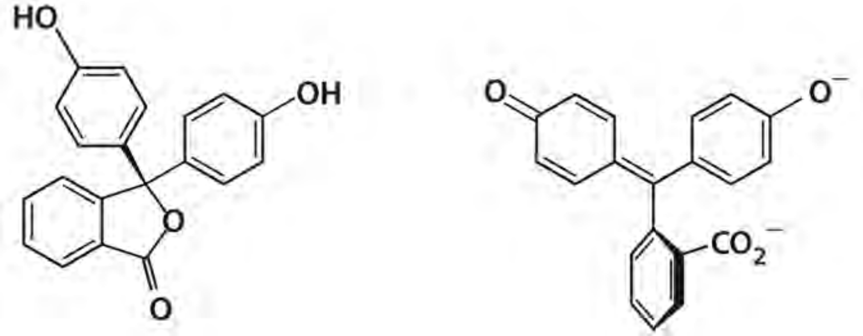


Delocalized pi bond that contains a total of 6 electrons, 1 from each carbon atom

In Figure 7.10, the sigma bonds shown in Figure 7.9 are now shown as just lines. This is to make the delocalized pi bonds clear. Too much clutter otherwise.

Now, why in the world did we go through all that? One reason is that it gives you more practice thinking about bonding in molecules and what happens in hybrid orbitals. The other is that benzene is such a special molecule in organic chemistry, it's good to know something about its structure. Plus, it shows up everywhere, in all kinds of molecules. For example, there are many benzene rings in the two versions of the phenolphthalein molecule I showed you in Chapter 5. Those are repeated in Figure 7.11.

Figure 7.11



---

[1] After writing that, I realized how upper-crust academic it sounded. No, one *doesn't* have to ask this question. It's just something chemists and physicists, who are aware that symmetry governs much in the universe, would ask.

Figure 7.12

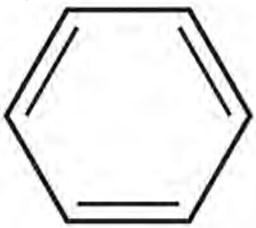
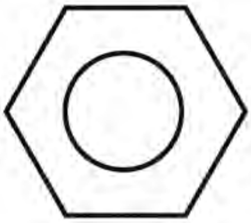


Figure 7.13



The benzene rings in those structures are represented as shown in Figure 7.12.

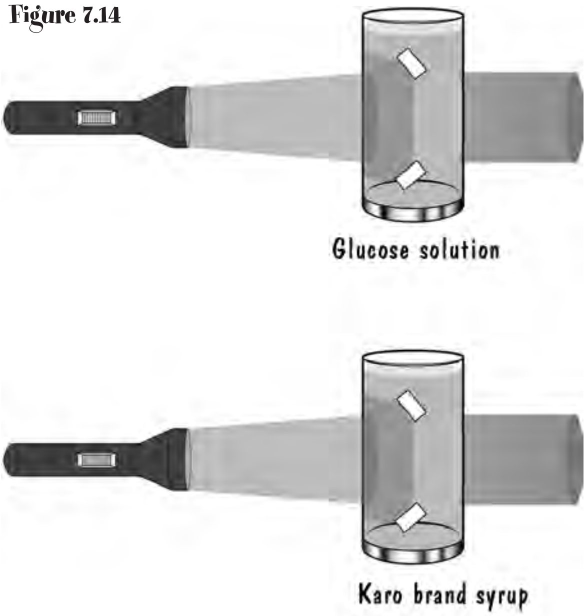
Figure 7.12 uses a bit of shorthand notation. First, each corner, or vertex, of the hexagon contains a carbon atom. It is also assumed that each corner has a hydrogen atom attached to it, as in Figure 7.5 (p. 125). So, any corner of a benzene ring in those phenolphthalein structures that doesn't have a line leading away to some other atom or molecule is assumed to have a hydrogen atom attached to it. Next notice the lines on the inside of the hexagon that seem to indicate double bonds. This is an old notation that hangs around today, and it's supposed to signify that there are both double and single bonds connecting the carbons in the hexagon. As we found out prior to this, though, in addition to the single sigma bonds connecting the carbons, there is one delocalized pi bond that contains six electrons. Thus, a better way to represent the benzene ring is shown in Figure 7.13.

More things to do before you read more science stuff

Get two identical clear glasses, a flashlight, about 10 grams of glucose powder dissolved in a glass of distilled water (this is the same glucose you used in Chapter 4), and a bottle of light corn syrup. (Karo brand works fine—make sure it's the clear stuff.) Finally, you need to find three polarizing filters. If you happen to have a copy of *Stop Faking It! Light*, then you have those filters in the pocket in the back of the book. If you don't have that book, you can obtain polarizing filters relatively cheaply from NSTA Press or a company called Rainbow Symphony.[2]

Once you have all the materials, set things up as shown in Figure 7.14. Tape the polarizing filters to the glasses.

Figure 7.14

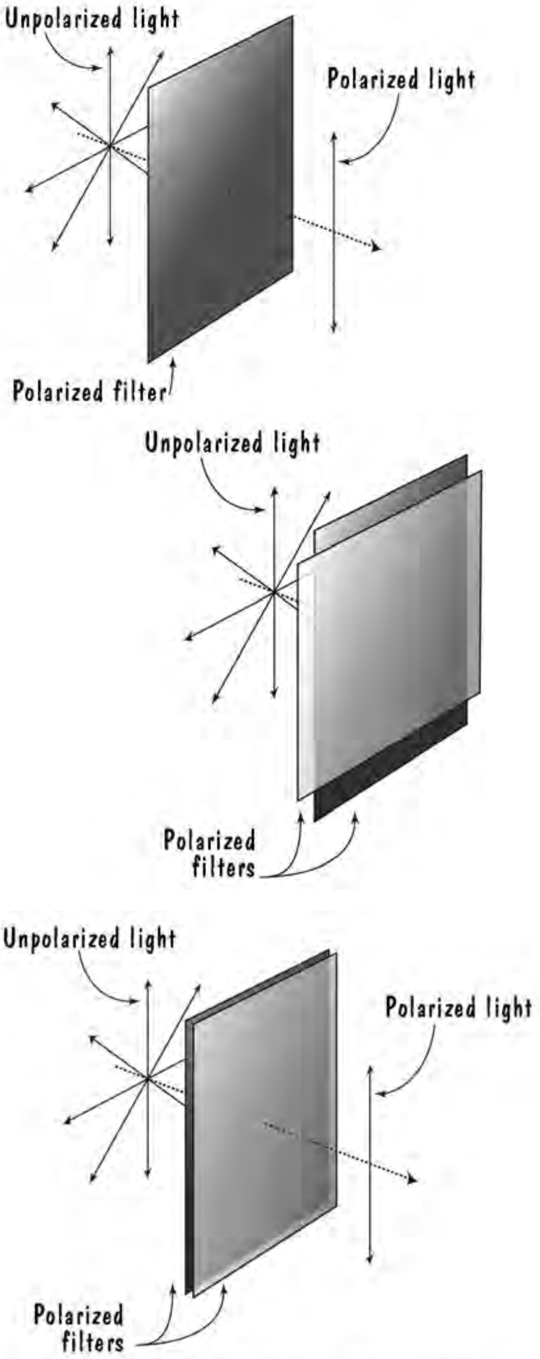


---

[2] You can get individual packets of polarizing filters (the packet also contains colored filters and something called a diffraction grating) by going to *store.nsta.org* and searching for PB169X3L, and you can reach Rainbow Symphony at *www.rainbowsymphony.com*.

Look through the third polarizing filter (this is the filter that's not taped to either glass) at the light emerging from the light corn syrup. Rotate the filter as you look through it and notice the pretty colors. Do the same with the glucose solution. Notice the pattern of colors—which colors occur in which order? If all goes correctly, you'll see the same pattern of colors when you rotate the filter to the left when viewing through the corn syrup and when you rotate the filter to the right when viewing through the glucose solution.

Figure 7.15



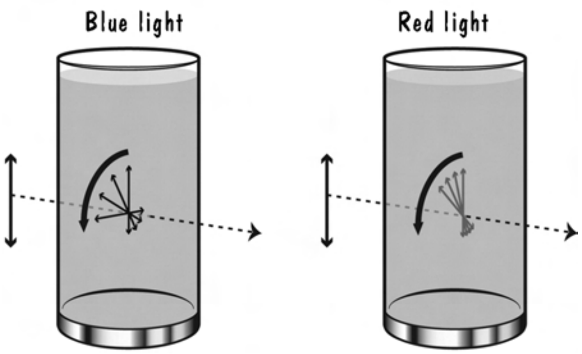
Depending on the orientation of the second filter, the light exiting the first filter does or doesn't get through the second filter.

More science stuff

I'd better start by explaining what polarizing filters do. One filter removes all light from the incoming light except that which is oscillating in a single direction. A second filter can let all, some, or none of that single-direction light (called **polarized light**) through, depending on the orientation of the two filters. Figure 7.15 shows how this works. The arrows indicate the direction of oscillation of the light waves.

Now let's apply our knowledge of polarizing filters to the corn syrup. When polarized light goes through corn syrup, the corn syrup rotates the direction of polarization. It also rotates different colors of light through different angles. Because the polarizing filter you have in your hand only lets light polarized in a given direction through, you will see different colors of the incoming light at different angles. As you rotate your filter to the left, you should see red, orange, green, and blue in that order. See Figure 7.16.

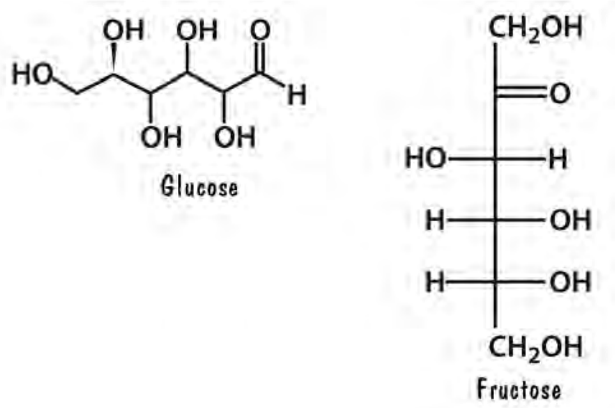
Figure 7.16



With glucose, the rotation of the polarized light is in the opposite direction—to the right as you look at the light coming at you. So as you rotate the filter in your hand to the *right*, you will see the same order of colors—red, orange, green, blue.

Okay, neat, but so what? To answer that, we have to investigate the molecules we're dealing with. There are two kinds of sugars—fructose (in the corn syrup) and glucose.[3] They both have the molecular formula $C_{12}H_{22}O_{11}$, but the atoms in the molecules are arranged differently. See Figure 7.17.

Figure 7.17



Molecules that are composed of the exact same atoms but have those atoms arranged differently are called **isomers** (the Greek word *isos* means *equal*, and the Greek word *meros* means *part*), and fructose and glucose are **structural isomers**. What that has to do with polarized light is that fructose rotates polarized light to the left (as you are looking at the oncoming light) and glucose rotates polarized light to the right (as you are looking at the oncoming light). Reflecting this property, fructose and glucose are also known as levulose (the Latin word *laevus* means *left*) and dextrose (the Latin word *dexter* means *right*).

Since this isn't a book about syrup, there better be a bigger lesson here, and there is. Because carbon atoms can combine in so many different ways, the world is full of isomers, and the different arrangements and orientations of atoms lead to molecules with the same molecular formula having very different properties. In our example, fructose has a sweeter taste than glucose, which is why many products use high-fructose corn syrup. So, the study of organic chemistry involves

---

[3] This is a small lie. Corn syrup has both kinds of sugar in it, but much more fructose than glucose, so it behaves a lot like a solution of pure fructose.

not just the composition of different molecules and how they interact but also the three-dimensional structures of those molecules and how those structures affect interactions.

Even more things to do before you read even more science stuff

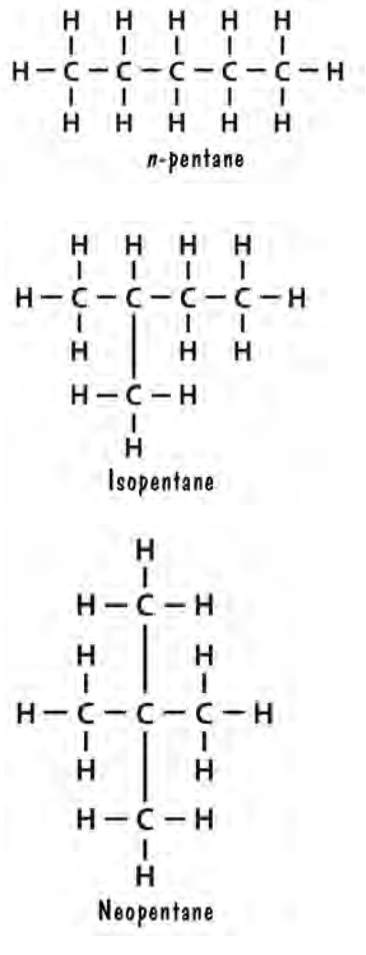
An activity section for those who don't want to be that active. All I want you to do is head to the internet or a library (that's a place with lots of books you can use for free—people used to use it a lot) and find just about anything do to with organic chemistry. Look at the names of the molecules, such as 1,2-propanediol, 1,2-benzenedicarboxylic anhydride, or 1,3-dimethylpentane. See if you can make sense of those. If you do any amount of organic chemistry, are you supposed to memorize those complicated names? Is there a pattern to follow? Why yes, there is, and I'll explain it in the next section.

Even more science stuff

So, how does one go about naming all these organic molecules? To address that, let's look at the organic molecule pentane. Its molecular formula is C_5H_{12}, and Figure 7.18 shows three different ways to put those atoms together.

The first structure, with all of the carbon atoms in a straight line, is called *normal pentane*, or *n*-pentane. The second structure, clearly different from the first, is called *isopentane*, and the third, different still, is called *neopentane* (named after the guy who saved the universe[4]). That's not too complicated, but remember that organic molecules can contain many, many carbon atoms. With 10 carbon atoms, decane ($C_{10}H_{22}$) has 75 different possible structures. Using a different prefix for each structure would be a bit cumbersome. So, chemists came up with a different way of naming organic molecules besides using prefixes. This system, known as the **IUPAC system** (IUPAC stands for International Union of Pure and Applied Chemistry), uses names for basic structures and then numbers to explain the different orientations. For example, the molecule CH_3 is known as the **methyl group**. Pentane is the molecule C_5H_{12}. The isomer of

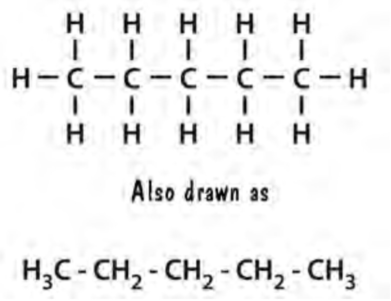
Figure 7.18



---

[4] Obscure movie reference. Yeah, I do that a lot.

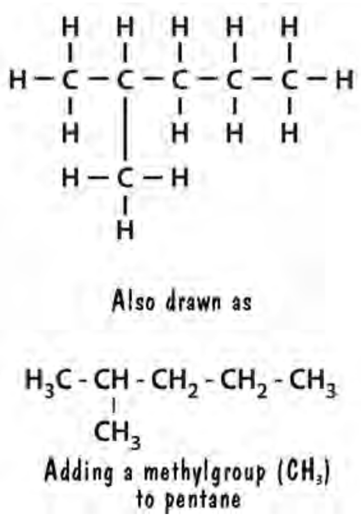
Figure 7.19



pentane in which all the carbon atoms are in a line is shown in Figure 7.19.

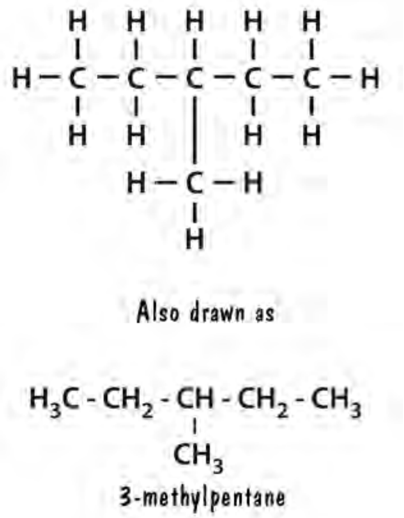
Figure 7.20 shows the pentane molecule with one of its hydrogen atoms replaced by a methyl group.

Figure 7.20



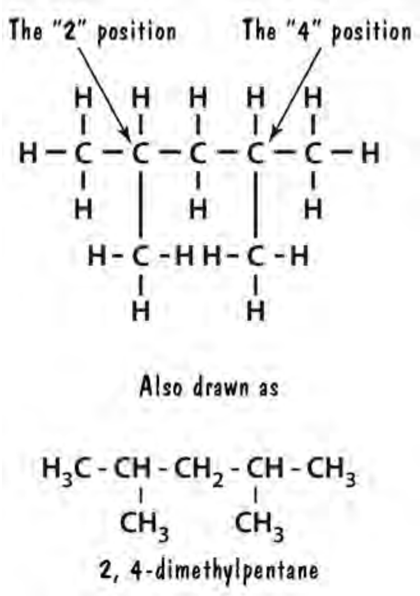
Now, in the prefix method of naming things, the molecule in Figure 6.12 (p. 119) would be called isohexane because there are six carbon atoms total and it's the first rearrangement you can create beyond having all six carbons in a row. In the IUPAC system, the molecule in Figure 6.12 is called 2-methylpentane. The number 2 indicates that the methyl group is attached to the second carbon after the start of the pentane chain. Similarly, the molecule in Figure 7.21 is called 3-methylpentane, because the methyl group is attached to the next carbon over in the chain.

Figure 7.21



We can add more than one methyl group to the pentane, as in Figure 7.22. The molecule shown there is called 2,4-dimethylpentane, with the numbers showing where the methyl groups attach and the *di* meaning there are two of them attached.

Figure 7.22



There's no way I can cover the complete naming system for organic compounds here, but maybe what I've presented will take the edge off when you see a molecule called 2-alkyl-4,4-dimethyl-2-oxazoline. Of course, scientists never make things easy, so you'll still see compounds called by their prefix naming convention. For example, we don't use the IUPAC system in naming isopropyl alcohol.

As in the first chemistry book, I've given you just a taste of the subject of organic chemistry. It's just plain impossible to be more comprehensive and maintain the purpose of the books in this series. Hopefully, though, what you have learned in this and the previous chemistry book in the series will help you feel a bit less intimidated if and when you tackle more traditional resources on organic chemistry.

Chapter summary

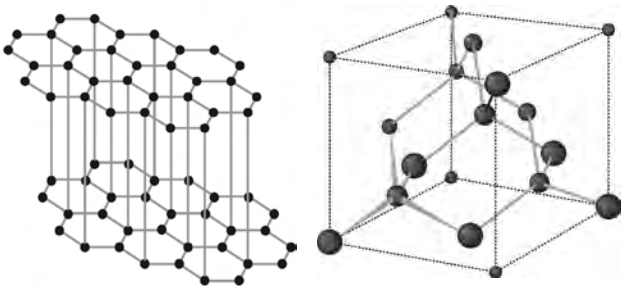
- Organic chemistry is the study of molecules that contain carbon and hydrogen, among other atoms. The molecules formed in organic chemistry usually depend on the hybridization of electron orbitals in carbon.
- Benzene is a ring of carbon atoms, with hydrogen atoms attached, that forms a unique kind of hybrid orbital with all the *p* orbitals. The result is a "delocalized" ring orbital in which all six carbon atoms donate one electron but share all six electrons in the ring.
- Carbon atoms often are explicitly omitted from many drawings of organic structures. Each vertex in such drawings is assumed to contain a carbon atom plus a hydrogen atom.
- The arrangement of atoms in an organic molecule can dramatically alter the properties of the molecule. Molecules that have the same atoms but different organizations of those atoms are called isomers.
- There are several methods for naming organic compounds. The most complete and unambiguous method is the IUPAC method, which uses numbers for the location of basic molecular groups attached to other basic molecular groups.

Applications

1. Most structural isomers have different boiling points and sometimes have quite different properties. Ethyl alcohol and methyl ether (isomers with the formula C_2H_6O) differ in their interactions with other molecules. Ethyl alcohol interacts violently with sodium metal, and methyl ether doesn't interact at all with sodium.
2. Two isomers familiar to everyone involve carbon atoms only. Graphite and diamond contain nothing but carbon atoms (if you ignore impurities) in different structures, which are shown in Figure 7.23 (p. 134).

   These different structures lead to vastly different properties (and costs!) of the materials. You can see in Figure 7.23 that graphite forms layers. These layers flake off easily, which is why we use graphite in pencils. Diamond, on the other hand, has a much more stable structure, a structure so stable that

Figure 7.23



diamond is one of the hardest known substances. Diamonds are also a lesson in the interaction of light with matter. Varying amounts of impurities (atoms that substitute for carbon in the structure) lead to different colors of diamonds. And to ask the question you're dying to ask, you *can* turn graphite into diamond. The intense pressure under the Earth's crust is best for this, but it can also be done commercially, leading to manufactured diamonds that are commonly used for industrial purposes such as drill bits.

SCILINKS THE WORLD'S A CLICK AWAY

Topic: Diamond and Graphite
Go to: *www.scilinks.org*
Code: MCB021

3. We commonly think of crude oil as the substance we convert to gasoline, but it's also the source of materials for thousands of products we use every day. Crude oil is the starting point for plastics, polyesters, nylon, and any number of other things, and it's all about converting some organic molecules into others. So, the availability of oil affects not only gas prices but also the prices of lots and lots of common products.

4. The representations of organic molecules shown in, for example, Figure 7.22 are somewhat limited. They imply a completely flat molecule, but in reality most organic molecules are three-dimensional in shape. You will run across drawings that attempt to correct this inadequacy, such as that shown in Figure 7.24.

Figure 7.24

Perspective drawings like this make it easier to determine the difference in structural isomers, for example. Of course, with three-dimensional computer modeling, you can get an even better picture of what's going on.

Glossary

absolute zero. Absolute zero is the lowest temperature on the Kelvin scale. It is theoretically impossible to achieve absolute zero due to Heisenberg's uncertainty principle.

acid. An atom or molecule that donates one or more protons in a chemical reaction. Also, something in vats you should stay away from if you're an extra in a horror movie.

activation energy. The energy input necessary to cause a chemical reaction to occur on its own. Spontaneous reactions have a negligible activation energy. Also, the energy input necessary to get my son up in the morning. This is a large amount of energy.

anode. The negative terminal on a battery, or the terminal that attracts positively charged ions in an electrochemical reaction. Also, the part of your face that's stuffed up when you have a code.

Arrhenius theory of acids and bases. A theory of acids and bases in which an acid is defined as any substance that contributes protons or, more accurately, hydronium ions, in solution; and a base is defined as any substance that contributes a hydroxide ion (OH^-) in solution.

atmosphere. A unit of pressure. which is roughly the atmospheric pressure at sea level. Also, something generally lacking in fast food restaurants.

atomic mass. The mass of a single atom of an element, measured in atomic mass units. Also, where atoms go on Saturday night.

atomic mass unit. A unit of mass based on the mass of carbon-12 being exactly 12.0. With this system, the mass of each atom is approximately equal to its atomic number. Also, a military hospital for atoms. Think about it.

atomic number. The number of protons in an atom/element. For neutral atoms, the atomic number is also equal to the number of electrons in the atom.

atomic theory. The theory that all matter is composed of tiny things called atoms. The theory includes not just the atoms themselves, but what they're made of and how they interact with other atoms. The theory seems to be working pretty well.

atomic wedgie. Something you want to avoid in your school years. This has little to nothing to do with atoms.

Avogadro's number. The number 6.02214×10^{23}. This is the number of atoms or molecules contained in a mole. Also, what was up when Avogadro died.

base. An atom or molecule that accepts protons in a chemical reaction. Also a molecule that produces hydroxide ions when dissolved in water. Also, a white canvas thing used in baseball. Also, how some people refer to my sense of humor.

benzene ring. A closed structure containing six carbon atoms with six hydrogen atoms attached. Also, what benzenes give to one another when they get married.

bioluminescence. The process in which living organisms produce their own light. Think glowing fish and fireflies.

Boltzmann's constant. A constant used in a number of physics and chemistry applications, equal to 1.380×10^{-23} joules/(degree Kelvin). In this book, Boltzmann's constant is present in one version of the ideal gas law.

bond. A connection between atoms that involves the transfer or sharing of electrons. Also, something you can do with baby humans and baby ducks provided you want them following you around a lot.

Brönsted-Lowry theory of acids and bases. A theory that labels as an acid anything that donates one or more protons in a chemical reaction and labels as a base anything that accepts one or more protons in a chemical reaction.

carbon dating. A process in which scientists use the radioactive decay of carbon isotopes to determine the age of a substance. Also, when carbon atoms take in dinner and a movie.

catalyst. A molecule that facilitates chemical reactions between other molecules. Enzymes are catalysts.

catalytic converter. A device on automobiles that uses catalysts to convert unburned hydrocarbons into carbon dioxide and water.

catalyzed reaction. A chemical reaction that uses a catalyst to proceed in a certain direction.

cathode. The positive terminal on a battery, or the terminal that attracts negatively charged ions in an electrochemical reaction.

chain reaction. A process in which one nuclear reaction triggers two or more other nuclear reactions, leading rapidly to a whole bunch of nuclear reactions happening at once.

chemical equation. A statement using chemical symbols that represents a chemical reaction, with reactants on the left, followed by an arrow or two, followed by products on the right. A balanced chemical equation has equal numbers of each kind of atom on either side of the arrow(s), and can tell you how much of one thing reacts with how much of another thing to produce how much of a new thing or two or three.

chemical reaction. Any interaction between atoms and/or molecules in which molecules change structure with the addition, subtraction, or substitution of various atoms. Many textbooks make a big deal out of distinguishing between a chemical change and a physical change, but personally I don't see much value in it.

cloud chamber. A place where clouds meet. Not really. It's a device that allows you to see tracks left by charged particles.

combustion. A chemical reaction in which hydrocarbons combine with oxygen to produce carbon dioxide and water. We usually call this burning.

concentration. A number that tells you how much of a chemical is dissolved in a solution. Concentration is typically expressed as a molarity. Also, a game show I used to watch as a kid, which, now that I think of it, wasn't that great. Hugh Downs was the host.

covalent bond. A bond in which atoms share one or more electrons in a hybrid orbital.

cyalume. A special chemical involved in the glowing of glow sticks.

delocalized orbital. A hybrid orbital containing electrons that are shared among a number of atoms, across a fairly large region of space, atomwise. Also, an orbital that is "not from around here."

dissociated. A description of what happens when ionically bonded molecules dissolve in water or some other solution.

electrochemistry. A term that covers all sorts of chemical reactions that involve the movement of ions or electrons from one place to another, either in solutions or wires. Also, chemistry done by electros.

electron. A negatively charged object that is part of an atom. Electrons are really tiny, so tiny in fact that they are considered point objects, with no size at all.

electron spin. A property of electrons that can have a value of + 1/2 or −1/2, also referred to as up or down spin. This is a quantum mechanical construct that isn't associated with electrons actually spinning. Also, what electrons do when trying to promote a certain political cause.

electronegativity. A number that tells how much an atom tends to attract electrons. Also, something not encouraged in electron motivational seminars.

electroplating. The process of using a separation of charge or a voltage difference to deposit a film of metal on a given object. This can be done in a solution or in a vacuum. Also, what they do on *Top Chef* when they're serving electros.

energy level. An allowed energy for electrons in an atom or a molecule. Also, what you use to keep energies straight when hanging them on your wall.

energy shell. Another name for an energy level. Also, what energies pull into when scared.

Erlenmeyer flask. A funny-shaped flask used in chemistry. The narrow neck prevents spills when you're swirling the contents around. Clearly invented by a guy named Erlenmeyer.

exponential decay. A reduction in the number of any kind of thing in which the number of items that disappear is a function of the number of items that are present. Radioactive decay is exponential, and all forms of radioactive decay can be characterized by a half-life. Also, what apparently happens to your teeth if you use methamphetamine. My teeth are fine, thank you.

exponential growth. A growth in the number of any kind of thing in which the number of new items produced is a function of the number of items that already exists. Exponential growth generally starts slowly and eventually increases at a rapid rate.

fission. A nuclear reaction in which a relatively large, unstable nucleus breaks apart into smaller nuclei and/or particles, usually with the help of bombardment by other particles. This results in the release of a great deal of energy. Fission is the source of energy for nuclear power plants and for certain kinds of atomic bombs.

fluorescence. The process in which electrons in molecules are excited to high energy levels, usually by the input of ultraviolet light, and then emit visible light after losing a bit of energy through vibration. This is the process behind day-glo paints and posters, and is part of what happens in glo sticks.

force. Any push, pull, nudge, whack, or other such thing that changes the motion of an object. Yoda extolls the virtues of this.

fusion. A nuclear reaction in which relatively small atomic nuclei and/or particles come together to form larger nuclei. This results in the release of a great deal of energy. Fusion is the source of energy for the Sun and for hydrogen bombs. Also, a form of jazz.

gram molecular mass. The mass of a molecule in atomic mass units. One gram molecular mass of a substance contains one mole of molecules of that substance. This would also be a great name for an important British scientist.

half-life. The time it takes for half of the existing amount of a radioactive substance to decay. Each radioactive element has a characteristic half-life. Also, what you've reached at the age of 50, provided you're going to live a long time.

half reaction. The oxidation or the reduction part of an oxidation-reduction reaction. Half reactions focus solely on the gain and loss of electrons, and ignore substances that aren't oxidized or reduced even though those substances might be part of the overall chemical reaction.

Heisenberg's uncertainty principle. A mathematical statement regarding the limits of our simultaneous knowledge of an object's position and momentum. There is also a version of this principle involving our knowledge of the energy of a process and the time for the process to occur. For the purposes of the content of this book, this principle states that we are limited in our knowledge of what electrons are doing and where they are.

hybrid bubble. A new, different-shaped bubble that forms when two or more regular soap bubbles come together. This concept doesn't make much sense except when used as an analogy for what happens in hybrid orbitals in atoms and molecules. Also, a bubble on a Prius.

hybrid orbital. An electron orbital that is the mathematical combination of two or more separate electron orbitals. A simple covalent bond is an example of a hybrid orbital, and electron orbitals within a single atom can also form hybrid orbitals. Also, an orbital favored by electrons concerned about the environment.

hydrogen burning. The process in which hydrogen undergoes fusion to produce helium. We usually use this term in describing what happens inside stars, including our Sun.

hydronium ion. An ion produced when a free proton combines with a water molecule. Its chemical symbol is H_3O^+.

ideal gas constant. Also known as the universal gas constant, it's a constant number, represented by R, contained in the ideal gas law. The value of R is 0.0821 (L-atom)/(mol-K).

ideal gas law. A relationship between pressure, volume, and temperature for an ideal gas, stated $PV = nRT$. An ideal gas is one in which we ignore any energy lost in collisions of molecules and ignore the vibration and rotation of molecules in the gas. Also, the gas law that all other gas laws look up to.

inert gases. Another name for noble gases, the name deriving from the fact that inert gases have filled outer shells of electrons and thus do not readily interact with other atoms.

ion. An atom or molecule with a deficiency of or excess of electrons. Ions are either positively or negatively charged, and can carry an electric current when in solution. Also, a very, very long time for bad spellers.

ionic bond. A bond between atoms in which one atom grabs electrons from another atom, resulting in an electric attraction. For bad spellers, a bond involving a result that is the opposite of what is expected and usually amusing.

isomer. A molecule that contains the same exact atoms as another molecule, but in a different structural arrangement.

isotope. An atom that contains the same number of protons, but a different number of neutrons, as another atom. In other words, a more or less massive version of the same element. Also, the team name for a minor league baseball team in Albuquerque, which was stolen from the TV show *The Simpsons.*

IUPAC system. A system for naming organic compounds. IUPAC stands for International Union of Pure and Applied Chemistry.

kinetic energy. The energy an object has because of its motion. Kinetic energy is equal to $\frac{1}{2}mv^2$.

kinetic theory of gases. A model of the behavior of gases that makes a number of simplifying assumptions about energy involved in collisions and the independence of gas molecules.

law of conservation of mass. A law stating that in a chemical reaction, no mass is gained or lost. This law holds true as long as you don't concern yourself with extremely small changes in mass. In actuality, mass is converted to energy or vice versa in every chemical reaction, but these changes are negligible in most chemical applications.

LED. A light emitting diode, which is an electrical device that emits light using a small electric current.

logarithmic scale. A mathematical scale used to compare very large and very small numbers on the same graph. A change of 1 unit in a logarithmic scale corresponds to an actual change of a factor of 10. pH is a logarithmic scale. Also, a device used to determine the mass of logarithms.

magnetic field. A model used to explain magnetic interactions. In this book, magnetic fields are used only as an example of a model and aren't important for understanding the chemistry.

mass. A numerical measure of an object's inertia, which tells how difficult it is to change an object's motion with an unbalanced force. Also, a kind of hysteria.

methyl group. One carbon atom combined with three hydrogen atoms. Also, a therapy session for methyls.

model. In science, a construct that helps explain observations. Scientific models can be physical or mathematical. Also, someone who makes a lot of money, especially the "super" variety.

molar. A unit of concentration. Something that has a concentration of 1.0 molar has one mole of substance per one liter of solution. Also, a tooth in the back of your mouth.

molarity. The concentration of a substance dissolved in solution, measured in moles per liter.

mole. A quantity of atoms, molecules, or in fact anything else that is equal to 6.02×10^{23} of the thing. Also, a bad TV show that I think is no longer on the air. Also, a kind of rodent. Also, something you should get checked out by a doctor if it changes in appearance. Also, if you add an accent to this word, it's a great Mexican sauce.

neutron. An uncharged particle that usually hangs around the nucleus of an atom, but is sometimes found roaming around on its own, contributing to chain reactions. Also, the last name of a pretty funny cartoon character.

noble gas. Any of the elements on the far right side of the Periodic Table, which are all gases at room temperature and are all reluctant to interact with other elements. Also ... Nah, we got away with the cartoon in the first chapter, so I'm not going to push my luck with the editors.

nonpolar covalent bond. A covalent bond in which the charge is evenly distributed, so that no single part of the resulting molecule is charged positively or negatively.

nucleus. The center of an atom, containing protons and neutrons (with the exception of hydrogen, which doesn't contain any neutrons). The nucleus doesn't take up much space in an atom. The word is pronounced "nu-clee-us," not "nu-cu-lus." The wrong pronunciation is like fingernails on a chalkboard to me, so naturally my son loves to say it incorrectly around me. Mispronunciation of the word knows no intellectual, economic, or cultural bounds.

orbital. A mathematical probability distribution that describes where one is likely to find one or more electrons in an atom or molecule.

organic chemistry. The study of chemical reactions that involve primarily molecules containing carbon and hydrogen, along with other atoms.

organic molecule. A molecule that contains primarily carbon and hydrogen. Also, a molecule that likes granola.

oxidation. A process in which an atom or molecule loses electrons.

oxidation-reduction (redox) reaction. A chemical reaction in which one or more atoms or molecules lose electrons and one or more atoms or molecules gain electrons.

oxidation states. Numbers assigned to atoms or molecules that help one keep track of the gain and loss of electrons in a redox reaction.

Periodic Table. A big ol′ chart that lists all the known elements along with information about each element. Many a middle school class has earned a pizza party when a certain percentage of students successfully memorize portions of this table. What a waste of pizza motivation power.

pH. A logarithmic scale used to measure the concentration of acid or base in a solution. When accompanied by the word *balanced,* this is a great selling point for shampoos, although most are pH balanced anyway.

phonon. A sound particle, analogous to a photon for light. Also, what people have when they′re texting.

phosphorescence. A process in which electrons in molecules get excited to higher energies with incoming light and then fall down to lower energies over a long period of time, emitting light of their own. This is what′s happening with glow-in-the-dark stickers and such.

photon. A light particle. This is a description of light that′s an alternative to waves. Also, the name of a torpedo on *Star Trek.*

pi bond. A ″sideways″ covalent bond that forms between two *p* orbitals from separate atoms. Pi bonds are generally weaker than sigma bonds. Also, something that matures over time when you invest in pies.

polar. A description of molecules in which there is an overall separation of positive and negative charges. Also, a type of bear.

polar covalent bond. A covalent bond in which electrons spend more time on one side of the bond than on the other, leading to an overall separation of positive and negative charges. Also, a covalent bond formed between two white bears.

polarized light. Light in which the light′s electric field vibrates in a specific direction. This is what you get when sending light through a polarizing filter. Many sunglasses are polarized because polarized filters eliminate roughly half of the light that′s reflected from a surface.

polymers. Long chains of identical organic structures. These are essential ingredients in bad fashion statements.

pressure. Force divided by the area over which the force is exerted. Also, what a writer feels when he′s extremely late in meeting a book deadline, which is what happened with this book.

probability distribution. A mathematical description of the probability of finding something in a given region of space. In chemistry, this applies primarily to electrons.

product. What's produced in a chemical reaction. Also, what you put in your hair if you're a metrosexual. Learned that term by watching *Queer Eye for the Straight Guy.*

proton. A positively charged particle that usually resides in the nucleus of atoms, though it can exist on its own.

radioactive decay. The process in which an element turns into another element with the release of various subatomic particles and energy.

radioactivity. Basically the same thing as radioactive decay.

reactant. A molecule or atom that reacts with another molecule or atom in a chemical reaction. Reactants are on the left side of a chemical equation.

reduction. A process in which an atom or molecule gains electrons. Also, a fancy name for gravy.

sea of electrons. A description of what electrons do in metals. Valence electrons in metals are shared among all the atoms in the metal.

shells. A name given to the different energy levels occupied by electrons in atoms and molecules. Also, something that adorns all sorts of household items such as lamps and toilet seats after a family has visited Florida.

sigma bond. A direct "head on" bond between various combinations of *s* and *p* orbitals from separate atoms.

solution. A chemical dissolved in a liquid, which is often water but can be other liquids including alcohol. Also, something you're looking for in your math homework.

s p d f. The labels for various electron orbitals.

speed. The distance traveled by something divided by the time it takes to travel that distance.

state. This can refer to many different situations in which a substance finds itself. In this book, it refers to solid, liquid, and gas as states of matter.

strong nuclear force. An extremely short range force that keeps the components of a nucleus together.

structural isomers. Isomers in which the arrangement of atoms in the molecules differs.

temperature. Something that tells you how hot or cold a substance is. More technically, it's a measure of the average kinetic energy of the molecules in a substance.

thermodynamics. The study of all things having to do with heat, temperature, thermal energy, and similar things. These quantities are often explained in terms of what various molecules are doing, which explains the presence of "dynamics" in the word.

thermoluminescence. A process in which electrons in molecules gain energy through the absorption of heat followed by those electrons losing energy and emitting light.

titration. A process in which one combines one solution with another, slowly adding them together until an "end point" is reached, often marked by the change in color of an indicator.

triboluminescence. A process in which electrons in molecules gain energy through mechanical means such as grinding things together, followed by those electrons losing energy and emitting light. Also, light produced by ancient cultures.

triple bond. A bond between atoms in which three separate pairs of orbitals combine for the purpose of sharing electrons. Triple bonds tend to be strong.

valence electron. An electron in the "outer shell" of an atom. This label is usually reserved for *s* and *p* electrons.

velocity. The speed of something with the addition of a direction in which the something is traveling.

volume. A specified amount of three-dimensional space. Also, a setting on an iPod that is usually way too high where my son is concerned.

Index

Page numbers in **boldface** type refer to figures.

A

Absolute zero, 38–39
Acetic acid, 86, 94
Acid(s), 70
  concentration of, 76–77, 83
  definition of, 73–75, 83
  molecular formulas of, 86
  pH testing of, 70–72, **72**
  strength of, 76–77, 83
  theories of, 73–75
Acid-base indicators, 97, 99–101, 105
Acid-base reactions, 78–82
Activation energy, **21,** 21–22, 69, **69,** 70, 83
Air, 3, **3**
Air pressure, 27
Alpha particle, 110, 117, **118**
Ammonia, **64,** 64–65, 70, 71, 76
Anodes, 96, **97**
Antacids, 86
Anti-neutrinos, 110
Argon, 15
Arrhenius, Svante, 73
Arrhenius theory of acids and bases, 73–75
Atmosphere (atm), 27, 40
Atom(s), 2–14
  definition of, 3
  energy levels of, 8–10
  in gases, 25, 27
  light production by, 10
  in liquids, 25
  mass of, 5
  nucleus of, 4
  number of electrons in, **4,** 13, 14
  number of neutrons in, **4,** 115
  number of protons in, 4, 13, 14, 110, 113
  positive and negative charges of, 4
  size of, 5
  in solids, 25
Atomic mass, 13–14, 36, 110
Atomic models, **2,** 2–5, **55,** 55–56
Atomic number, 110, 113
Atomic theory, definition of, 2
Atomic units, 13
Avogadro's number, 35
Avogadro's principle, 33–36

B

Baking soda
  pH of, 70
  reaction with vinegar, 20–21
  use as antacid, 86
Bases, 70
  concentration of, 76–77, 83
  definition of, 73–75, 83
  molecular formulas of, 86
  pH testing of, 70–72, **72**
  strength of, 76–77, 83
  theories of, 73–75
Batteries, **90,** 90–96, **91, 92, 93, 95, 96, 97,** 104, 105
Benzene
  carbon bonding in, 124–125, 126
  ring structure of, 126, 127–128, **128,** 133

Beryllium, 11, 15
Beta particles, 121–122
Bioluminescence, 106
Bleach, 70, 71
Boltzmann's constant, 38
Bonds, 16–20, **17**
  covalent, 17–18, **18,** 19, 56, 63
  double, 57, 66, 126, 128
  pi, 57, **57,** 63, 66, **126,** 126–127, **127**
  sigma, 51, **51,** 53, **53,** 57, 58, 63, 124, 126, **126, 127**
  triple, 57, **57,** 66
Boron, 15
Boyle, Robert, 33
Boyle's law, 33–34
Bromine, 16
Bromothymol blue, 97, 98, 100–101
Brönsted-Lowry theory, of acids and bases, 75
Bubbles, **50,** 52, 53
  hybrid, **50,** 52

C

Calcium carbonate, 78
Calcium hydroxide, 70–71, 78–79
Carbon, 22
  hybrid orbitals of, 64
  isomers of, 130–133
  molecular shape of, 61
  valence electrons of, 59, **60**
Carbon dating, 121
Carbon dioxide, 83–84, 86
Catalysts, 68, 70, 83
Catalytic converters, 83–86, **84, 85**
Cathode, 96, **97**
Cerium oxide, 70, **70**
Chain reactions, 119
Charles, Jacques, 33
Charles' law, 33–34
Chemical equations, 20
  balanced, 21, 41, 78, 84, 87
Chlorine, 16
Cloud chamber, 108–109, **109,** 110
Color
  absorption and reflection of, 104
  changes in chemical reactions, 97–101
  of light, 97–101, 128–131
Combustion, 68–70
Combustion reactions, 69–70, 83–84
Copper, electrochemistry of, 92, 95, 96, **97**
Copper sulfate, electrochemistry of, 90–91, 93, **93**
Corrosion, 105
Covalent bonds, 17–18, **18,** 56, 63
  nonpolar, 18, 19
  polar, 18–19
Crude oil, 134
Cubes, 51, **52,** 58
Cyalume, 103, **104**

D

Deuterium, 117–118
Diamonds, 133–134, **134**
Diode, 91
Dissociation, of molecules, 73
Double bonds, 57, 66, 124, 126, 128
Dyes, radioactive, 113, 120–121

E

Earth, 3, **3**
Einstein's theory of special relativity, 116
Electrical circuits, 90–97
Electrical conduction, 90–97, 104
Electric current, 92–93, 94, 95, 104
Electricity, 92
Electric repulsion, 55, **55,** 62–63, **63,** 73, **73,** 118
Electrochemistry, 89–97, 104, 105–106

Electro-luminescence, 89–106
  electrochemistry, 89–97, 104, 105–106
  light-matter interactions, 89, 97–106
Electrolytes, 96, 104, 105
Electromagnetic waves, 6, **6,** 10
  light waves, 6
Electron(s)
  charge on, 4, 8
  cloud of, 5
  energy levels of, 5–6, **6, 7,** 14, 100
  interaction with light, 99, 104
  interaction with photons, 46, **47**
  mass of, 5
  in metals, 20
  negative charge on, 55
  number in atoms, **4,** 13, 14
  orbitals of, 7, 8–9, **9**
  probability distribution of, 7, **7, 8,** 47
  sea of, 20
  size of, 5
  valence, 14–15, 16, 18, **18**
  wave function of, 48
Electronegativity, 79
Electron shells
  filled and unfilled, 10
  inner, 14, 15
  outer, 14, 15, 17–18
Electron spin, 54–55, **55, 60**
Electroplating, 93, **93,** 104
Elements, 11
  main, 2, **3**
  Periodic Table of, **11,** 11–14, **12**
E = mc^2, 116
Energy
  activation, **21,** 21–22, 69, **69,** 70, 83
  of chemical reactions, 1–3
  equals mass, 116–117, 118, 120
  in radioactivity, 119–120
Energy levels, 8, 51, 52, **52**
  of atoms, 8–9
  diagram for oxygen, **9,** 9–10
  of electrons, 5–6, **6, 7,** 14, 100
  of hybrid orbitals, 59
  and patterns of elements in Periodic Table, 14–16
  in ultraviolet light, 101–103
Energy shells, 5, 116
  of neutrons, 120
  of protons, 120
Erlenmeyer flasks, 79, **79,** 80
Ethene, 125, **126**
Ethyl alcohol, 133
Exponential decay, 111, **111,** 112

F

Fire, 3, **3,** 68, **68**
Fission, 119, 120
Fluorescence, 102–103, 105, 106
Fluorine, 16
Force, 23, 24, 26
Fructose, 130, **130**
Fusion, 118–119, 120

G

Gamma particles, 110, 122
Gas(es), 24, 25
  ideal, 25–28, **26,** 30–32, 38
  inert, 14–16
  kinetic energy of, 29–30
  molecular measurement of, 25–26
  moles of, 35–36, 40–41
  number of molecules in, 25–26, 33–38, 40–41
  pressure of, 26, 27, 28, 30, 31, 32, **32,** 38
  speed of, 23–24, 28, 30, **30**
  temperature of, 28, 29, 30, **30,** 31, 32, **32,** 38, 39
  volume of, 26, 27, 28, 31, **31,** 33, 42
Gas laws, 32–34, 35–36

Gay-Lussac, Joseph-Louis, 32
Gay-Lussac's law, 33–34
Glucose
  molecular structure of, 130, **130**
  oxidation of, 81
  positive and negative charges on, 81
  reaction with hydroxide ions, 79–82
  as structural isomer, 130
Glucuronic acid, 80–82
Gram molecular mass, 37
Graphite, 133–134, **134**
Grik, 35

H

Half-life, 111–113
Half-reactions, 94–95, 104
Heisenberg uncertainty principle, 47–48, 49
Helium, 14, 28
  electron spin in, **55**
  kinetic energy of, **30**
  measurement units of, **26,** 27
  nucleus of, 110–111, 118, **118**
  number on Periodic Table, 11
  valence electrons of, 14–15
Hybrid bubbles, **50,** 52
Hybrid orbitals, 51, 53, **56,** 56–57, **57,** **58,** 58–63, 64–65
Hydrocarbons, 83–84, 84, 85, 86
Hydrochloric acid, 70–71, 72
  molarity of, 76
  pH of, 76
  reaction with zinc, 41–42
Hydrogen, 4, 18
  atomic mass of, 13, 14
  bonding with oxygen, **18,** 18–19
  number on Periodic Table, 11, 13
  s orbitals of, 51, **51**
  valence electrons of, 14–15
Hydrogen chloride, 73, **74**
Hydrogen ions, concentration of, 73–74, 75–76
Hydrogen peroxide, 103, **104**
Hydronium ions, concentration of, 74, 75, 76, 83
Hydroxide ions, 74
  reaction with glucose, 79–82

I

Ice, 24
Ideal gases, 25–28, **26,** 30–32, 38
Ideal gas laws, 34–35, 37–38, 39–42
Inert gases, 14–16
Inverse relationship, 33
Iodine, 16
  radioactive, 120–121
Ionic bonds, 1**7,** 16–17
Ions, 13
Isomers, 130–134
Isotopes, 4, 13, 118–119
IUPAC system, 131–33

K

Kinetic energy, 23, 24, 29–30, **30,** 38, 116
Krypton, 15–16

L

Law of conservation of mass, 21
LED (light-emitting diode), 91, 92, 94
Lemon juice, pH of, 76
Levulose, 130
Light
  color of, 97–101, 128–131
  interaction with electrons, 99, 104
  interaction with matter, 89, 97–106
  photons in, 45–46
  polarized, **129,** 129–131, **130**
  production by atoms, 10

    rotation of, 128–131
    ultraviolet, 101–104
Light waves, 6
Liquids, 24, 25
    electric current conduction in, 90–95, 104
    pH of, 76
    pH testing of, 70–72
Lithium, 15
Litmus paper, 70, 71, 72
Logarithms (logs), 71–72, 75, 76

M

Magnet(s), 113–114, **116**
Magnetic fields, 2, **2**
Mass, 23, 24
    atomic, 5, 13–14
    of electrons, 5
    equals energy, 116–117, 118, 120
    in kinetic energy, 29
    Law of Conservation of, 21
    of neutrons, 5
    of nucleus, 5
    of protons, 5
Matter
    interaction with light, 89, 97–106
    states of, 23, 24
Measurement, in chemistry, 35–38
Metals
    corrosion of, 105
    electroplating of, 93, **93,** 104
    free electrons of, 20
Methane
    gram molecular mass of, 40
    number of molecules in, 40–41
    orbital bonds of, 124, **124, 125**
Methylene blue, 79, 80, 81
Methyl ether, 133
Methyl groups, 131–132, **132**
Models, scientific, **2,** 2–3
Molarity, 23, 37, 71, 75, 76–77, 98
Mole, 23, 27, 36
    of gases, 35, 36, 40–41
Molecular formulas, 86
    in acid-base reactions, 81
Molecular structure
    effect of electron orbitals on, 50–66
    effect of light on, 99
    effect on light emission, 104
Molecules, 16
    dissociation of, 73
    in gases, 25–26, 33–38, 40–41
    in liquids, 25
    in solids, 25
Momentum, 49
Motion energy, 29–30

N

Neon, 15
    kinetic energy of, **30**
    measurement units of, **26,** 27
Neutrinos, 110
Neutrons, 4
    energy shells of, 116, 120
    mass of, 5
    number in atoms, **4,** 115
    ratio to protons, 115, 116
Nitric oxide, 84
Nitrogen, orbitals of, 55–56
Nitrogen oxide, 83–84
Noble gases, 14–16
Nuclear chemistry, 107–122
Nucleons, 115
Nucleus
    charge on, 4
    mass of, 5

O

Objects, location of, **44,** 44–46, **45**
Observation, 47–48, 49, 63
Orbitals, 7, 8–9, **9,** 10, **10**

  in atomic bonds, 43
  covalent bonds of, 17–18, **18,** 19, 56–57, 63
  d, **48,** 66
  delocalized, 127, **127**
  double bonds of, 57, 66, 126, 128
  effect on molecular shape, 50–66
  filled, 55–56, 62–63, 64, 66
  "fuzzy," 43
  hybrid, 51, 53, **56,** 56–57, **57, 58,** 58–63, 64–65
  model of, **55**
  of organic molecules, 124
  p, **48, 53, 55–56,** 58–59, **59,** 125
  partially filled, 55
  pi bonds of, 57, **57,** 63, 66, **126,** 126–127, **127**
  probability distribution of, 47–48
  s, **48,** 51, **51,** 53, **53, 55–56,** 58–59, **59,** 125, **125**
  shape of, 43–44, 53–54
  sigma bonds of, 51, **51,** 53, **53,** 57, 58, 63, 124, 126, **126, 127**
  sp, 58–61, **61, 62,** 64, **64,** 124, 125, 126
  triple bonds of, 57, **57,** 66
  unfilled, 66
Organic chemistry, 22, 123–134
Organic molecules, 68
  names of, 131–133
  perspective drawings of, 134, **134**
Oxidation, 82, 85–87, 91
Oxidation reactions, 85
Oxidation-reduction reactions, 82, 83
  balanced chemical equations in, 87
  in batteries, 105
Oxidation states, 81, 85, 86, 87
Oxygen
  bonding with hydrogen, **18,** 18–19
  in catalyzed reaction with sucrose, 68–70
  charge in gluconic acid, 81
  energy level diagram for, **9,** 9–10
  hybrid orbitals of, 62, **62**
  liquid, 25
  number on Periodic Table, 11, 13
  valence electrons of, 62

P

Palladium, 84, 85, **85**
Pascals, 40
Pentane, 131, **131,** 132, **132**
Percentages, 41
Periodic Table, 11–14, **12**
  electronegativities of elements on, 17
  meanings of letters and numbers ion, 11, 13–14
  use in chemical formulas, 78
pH
  of acid-base reactions, 101
  definition of, 75–76, 83
  of the human body, 86
Phenol, 103, **104**
Phenolphthalein, 97–98, 99–100, 128
pH indicators, 105
Phonon, 46
Phosphorescence, 103, 105, 106
Photons
  interaction with electrons, 46
  in light, 45–46
  movement of, 46
pH paper, 70, 71, 72, 75, 76
pH scale, 75–76
pH testing, 70–72
pi, 43
pi bonds, 57, **57,** 63, 66, **126,** 126–127, **127**
Platinum, 84, **84, 85**
Plutonium, 120
Point objects, 5
Polarizing filters, **128,** 128–130, **129**
Polonium, 110, 117

Polyester, 22
Polymers, 22
Positrons, 110, 122
Potassium, 15–16
Pressure
  definition of, 26
  of gases, 26, 27, 28, 31, 32, **32,** 38
  measurement units of, 27, 40
  relationship with temperature, 27, 30, 32, **32,** 33, 38–40
  relationship with volume, 31, **32,** 33
Probability distribution, of electrons, 7, **7, 8,** 47
Products, of chemical reactions, 20–21
Protons
  acceptance by bases, 75
  charge on, 4
  donation by acids, 75
  energy shells of, 116, 120
  mass of, 5
  number in atoms, 4, 13, 14, 110, 114
  ratio to neutrons, 115, 116

Q
Quantities, measurement of, 35–38
Quantum mechanics, 5–6, 38–39, 47

R
Radiation, acceptable levels of, 113
Radioactive decay, 111–113, 120, 121–122
Radioactive elements, half-life of, 111–113
Radioactivity, 109–122
  energy in, 119–120
Radon, 110, 113, 117, **117**
Reactants, 20–21
Reactions, 20–22, 67–87
  acid-base, 78–82
  catalyzed, 68–70
  color changes in, 97–101
  combustion, 69–70, 83–84
  products and reactants of, 20–21
Redox reactions, 82, 83
  balanced equations in, 87
  in batteries, 105
Reduction reactions, 85
Relativity, 116
Rhodium, 84, **85,** 85

S
Safety concerns, ix
Salt
  as acid-base reaction product, 79, 83
  electrochemistry of, 90
  as reaction product, 79
Saltwater, electrical circuits in, 90–92
SciLinks
  acids and bases, 72
  atomic models, 3
  atomic orbitals, 43
  atomic theory, 2
  Avogadro's constant, 35
  carbon dating, 121
  diamond and graphite, 134
  electronegativity, 17
  electrons in an atom, 54
  fission/fusion, 119
  half-life, 108
  inert gases, 14
  kinetic energy, 29
  light and color, 97
  molecular shapes, 100
  Periodic Table, 13
  pH scale, 75
  radioactive decay process, 111
  redox reactions, 82
Shells, 5
  inner, 14

outer, 14, 17–18
Sigma bonds, 51, **51,** 53, **53,** 57, 58, 63, 124, 126, **126, 127**
Silver, 11
Small-scale systems, 46–48, 63
Sodium acetate, 20–21
Sodium bicarbonate, 20–21
Sodium chloride, interaction with water, 73, **73**
Sodium hydroxide, 70–71
as base, 74
proton acceptance by, 75
Solids, 24, 25
Sound, 46
Speed, 23, 24
comparison with velocity, 29–30
effect of temperature on, 28, 29, 30, **30**
of gases, 28, 29
Spheres, 51, **52**
Stadium Checkers game, 9, 14
Standard Temperature and Pressure (STP), 36–37
Steam, 24
Strong nuclear force, 115, 116, 118, 119, 120
Sucrose, in catalyzed reaction with oxygen, 68–70
Sulfur, gram molecular mass of, 37

T

Temperature
absolute zero, 38–39
Celsius, 27
definition of, 29, **30**
Fahrenheit, 27
of gases, 28, 29, 31, 32, **32,** 38, 39
Kelvin, 27, 38–39
relationship to kinetic energy, 38
relationship to pressure, 27, 30, 32, **32,** 33, 38, 39–40
relationship to speed, 30, **30**
relationship to velocity, 27
Thermodynamics, 23–42
Thermoluminescence, 105
Thermometer, 29
Triboluminescence, 105–106
Triple bonds, 57, **57,** 66
Tritium, 117–118

U

Ultraviolet light, 101–104
Uncertainty principle, 47–48, 49, 63, 64
Universe, 3, 8
Uranium, 118–119, 120

V

Valence electrons, 14–15, 16, 18, **18**
Velocity, 24, 27, **30**
Vinegar
electrochemistry of, 94, 96
reaction with ammonia, 79
reaction with baking soda, 20–21
Vinyl, 22
Volatile substances, 22
Volume
of gases, 26, 27, 28, 31, **31,** 33, 42
measurement of, 27
relationship to pressure, 31, **32,** 33
relationship to temperature, 38

W

Water, 3, **3,** 18, **18**
as acid-base reaction product, 83
as combustion reaction product, 83
freezing temperature of, 27, 36
gram molecular mass of, 37
polarity of, 81
Water molecule, 24–25

- molecular interactions of, **73,** 73–74
- polarity of, 72, 73, **73,** 104

Wave, resonant energy of, 6

Wave function, 48

Z

Zinc
- electrochemistry of, 94–95, 94–96
- half-reactions of, 94–95
- reaction with hydrochloric acid, 41–42