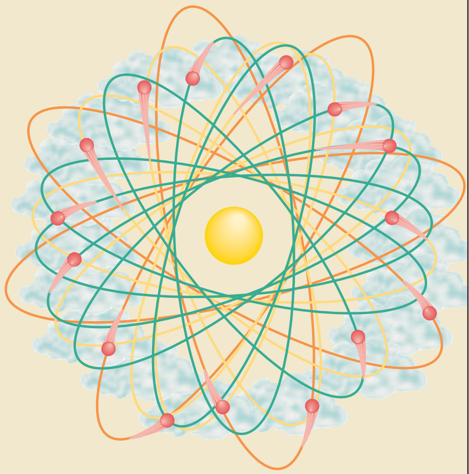


CHAPTER 2

Simple Things of Life

2

Chapter Outline

It is often helpful when learning new material to have the goals clearly stated before that material is presented. It is also helpful to have some idea why the material will be relevant. This information can provide a framework for organization as well as serve as a guide to identify the most important facts. The following table will help you identify the key topics of this chapter as well as the significance of mastering those topics.

Key Concepts	Applications
Understand that all matter is composed of atoms.	• Understand why you learn chemistry in a biology class.
Learn the basic structure of atoms.	• Understand the difference between atoms, elements, molecules, and compounds.
Learn what an isotope is.	• Understand how isotopes differ and how they are used.
Understand how to differentiate between diferent types of molecular bonds.	• Know how atoms stick together to form compounds.
Describe the chemical differences among acids, bases, and salts.	• Identify compounds that are acids, bases, or salts. • Work with the pH scale.
Understand the various states of matter.	• Describe the differences among liquids, solids, and gases.
Recognize that compounds may be broken down and reconnected in different ways.	• Understand that a chemical reaction is a recombining of atoms. • Know how to tell one type of reaction from another.
Understand how information is stored in the periodic table of the elements.	• Be able to use the periodic table of the elements to diagram various elements. • Understand the chemical and physical characteristics of various elements. • Be able to use this information to show how atoms may chemically bond.

PART TWO Cells Anatomy and Action

2.1 The Basics: Matter and Energy

In order to understand living things and how they carry out life's functions, you must understand what they are made of. All living things are composed of and use chemicals. There are more than 100,000 chemicals used by organisms for communication, defense, aggression, reproduction, and various other activities. For example, humans are composed of the following chemicals: oxygen (65%), carbon (18%), hydrogen (10%), nitrogen (3%), calcium (2%), and many others at lower percentages. Chemicals are also known as *matter*. **Matter** is anything that has mass and also takes up space (volume). Mass is how much matter there is in an object; weight refers to the amount of force with which that object is attracted by gravity. For example, a textbook is composed of the same amount of matter (its mass) whether you measure its mass on the Earth or on the Moon. However, because the force of gravity is greater on the Earth, the book will weigh more on Earth than if it were on the Moon. Both mass and volume depend on the amount of matter you are dealing with; the greater the amount, the greater its mass and volume, provided the temperature and pressure of the environment stays the same.

Two other features of matter are density and activity. **Density** is the weight of a certain volume of material; it is frequently expressed as grams per cubic centimeter. For example, a cubic centimeter of lead is very heavy in comparison to a cubic centimeter of aluminum. Lead has a higher density than aluminum. The activity of matter depends almost entirely on its composition.

All matter has a certain amount of energy, something an object has that enables it to do work or causes things to move. This chapter will focus on two types of energy, kinetic and potential. **Kinetic energy** is energy of motion. The energy an object has that can become kinetic energy is called **potential energy.** You might think of potential energy as stored energy. When we talk of chemical energy, we are really talking about potential energy in chemicals. This energy can be released as kinetic energy to do work such as moving chemicals to perform chemical reactions; that is, chemicals (matter) are broken apart and reassembled into other kinds of chemicals. An object that appears to be motionless does not necessarily lack energy. Its individual molecules will still be moving, but the object itself appears to be stationary. An object on top of a mountain may be motionless, but still may contain significant amounts of potential energy. Keep in mind that potential energy increases whenever things experiencing a repelling force are pushed together. You experience this every time you "click" your ballpoint pen and compress the spring. This gives it more potential energy that is converted into kinetic energy when the ink cartridge is retracted into the case. Potential energy also increases whenever things that attract each other are pulled apart. An example of this occurs when you stretch a rubber band. That increased potential energy is converted to the "snapping" back of the band when you let go. One of the important scientific laws, the *law of conservation of energy* or the **first law of thermodynamics,** states that energy is never created or destroyed. Energy can be converted from one form to another but the total energy remains constant. The amount of energy that a molecule has is related to how fast it moves. **Temperature** is a measure of this velocity or energy of motion. The higher the temperature, the faster the molecules are moving.

The three **states of matter**—solid, liquid, and gas—can be explained by thinking of the relative amounts of energy possessed by the molecules of each. A **solid** contains molecules packed tightly together. The molecules vibrate in place and are strongly attracted to each other. They are moving rapidly and constantly bump into each other. The amount of kinetic energy in a solid is less than that in a liquid of the same material. Solids have a fixed shape and volume under ordinary temperature and pressure conditions. A **liquid** has molecules still strongly attracted to each other but slightly farther apart. Because they are moving more rapidly, they sometimes slide past each other as they move. While liquids can change their shape under ordinary conditions, they maintain a fixed volume under ordinary temperature and pressure conditions; that is, a liquid of a certain volume will take the shape of the container into which it is poured. This gives liquids the ability to flow. Still more energetic are the molecules of a **gas.** The attraction the gas molecules have for each other is overcome by the speed with which the individual molecules move. Because they are moving the fastest, their collisions tend to push them farther apart, and so a gas expands to fill its container. The shape of the container and pressure determine the shape and volume of gases. A common example of a substance that displays the three states of matter is water. Ice, liquid water, and water vapor are all composed of the same chemical—H_2O. The molecules are moving at different speeds in each state because of the difference in kinetic energy. Considering the amount of energy in the molecules of each state of matter helps us explain changes such as freezing and melting. When a liquid becomes a solid, its molecules lose some of their energy; when it becomes a gas, its molecules gain energy.

All matter is composed of one or more types of substances called *elements*. **Elements** are the basic building blocks from which all things are made. Elements are units of matter that cannot be broken down into materials that are more simple by ordinary chemical reactions. You already know the names of some of these elements: oxygen, iron, aluminum, silver, carbon, and gold. The sidewalk, water, air, and your body are all composed of various types of elements combined or interacting with one another in various ways. The **periodic table of the elements** (How Science Works 2.1) lists all the elements. Don't worry, you will not have to know the entire table; only about 11 elements are dealt with in this text. The main elements comprising living things are C, H, O, P, K, I, N, S, Ca, Fe, Mg (i.e., C Hopkins Café, Mighty Good!).

Each single unit of a particular element is called an **atom.** Under certain circumstances atoms of elements join together during a chemical reaction to form units called

HOW SCIENCE WORKS 2.1

The Periodic Table of the Elements

Traditionally, elements are represented in a shorthand form by letters. For example, the symbol for water, H_2O, shows that a single molecule of water consists of two atoms of hydrogen and one atom of oxygen. These chemical symbols can be found on any periodic table of elements. Using the periodic table shown here, we can determine the number and position of the various parts of atoms. Notice that the atoms numbered 3, 11, 19, and so on, are in column IA. The atoms in this column act in a similar way because they all have one electron in their outermost layer. In the next column, Be, Mg, Ca, and so on, act alike because these metals have two electrons in their outermost electron layer. Similarly, atoms 9, 17, 35, and so on, have seven electrons in their outer layer.

Knowing how fluorine, chlorine, and bromine act, you can probably predict how iodine will act under similar conditions. At the far right in the last column, argon, neon, and so on, act alike. They all have eight electrons in their outer electron layer. Atoms with eight electrons in their outer electron layer seldom form bonds with other atoms.

Key: 1 = Atomic Number; Hydrogen = Name; H = Symbol; 1.0079 = Atomic Weight (Mass Number)

Representative Elements (s Series): IA, IIA. Transition Metals (d Series of Transition Elements): IIIB–IIB. Representative Elements (p Series): IIIA–VIIIA.

Period	IA	IIA	IIIB	IVB	VB	VIB	VIIB	VIIIB	VIIIB	VIIIB	IB	IIB	IIIA	IVA	VA	VIA	VIIA	VIIIA
1	1 Hydrogen H 1.0079																	2 Helium He 4.0026
2	3 Lithium Li 6.941	4 Beryllium Be 9.0122											5 Boron B 10.811	6 Carbon C 12.0112	7 Nitrogen N 14.0067	8 Oxygen O 15.9994	9 Fluorine F 18.9984	10 Neon Ne 20.179
3	11 Sodium Na 22.989	12 Magnesium Mg 24.305											13 Aluminum Al 26.9815	14 Silicon Si 28.086	15 Phosphorus P 30.9738	16 Sulfur S 32.064	17 Chlorine Cl 35.453	18 Argon Ar 39.948
4	19 Potassium K 39.098	20 Calcium Ca 40.08	21 Scandium Sc 44.956	22 Titanium Ti 47.90	23 Vanadium V 50.942	24 Chromium Cr 51.996	25 Manganese Mn 54.938	26 Iron Fe 55.847	27 Cobalt Co 58.933	28 Nickel Ni 58.71	29 Copper Cu 63.546	30 Zinc Zn 65.38	31 Gallium Ga 69.723	32 Germanium Ge 72.59	33 Arsenic As 74.922	34 Selenium Se 78.96	35 Bromine Br 79.904	36 Krypton Kr 83.80
5	37 Rubidium Rb 85.468	38 Strontium Sr 87.62	39 Yttrium Y 88.905	40 Zirconium Zr 91.22	41 Niobium Nb 92.906	42 Molybdenum Mo 95.94	43 Technetium Tc (99)	44 Ruthenium Ru 101.07	45 Rhodium Rh 102.905	46 Palladium Pd 106.4	47 Silver Ag 107.868	48 Cadmium Cd 112.40	49 Indium In 114.82	50 Tin Sn 118.69	51 Antimony Sb 121.75	52 Tellurium Te 127.60	53 Iodine I 126.904	54 Xenon Xe 131.30
6	55 Cesium Cs 132.905	56 Barium Ba 137.34	*57 Lanthanum La 138.91	72 Hafnium Hf 178.49	73 Tantalum Ta 180.948	74 Tungsten W 183.85	75 Rhenium Re 186.2	76 Osmium Os 190.2	77 Iridium Ir 192.2	78 Platinum Pt 195.09	79 Gold Au 196.967	80 Mercury Hg 200.59	81 Thallium Tl 204.37	82 Lead Pb 207.19	83 Bismuth Bi 208.980	84 Polonium Po (209)	85 Astatine At (210)	86 Radon Rn (222)
7	87 Francium Fr (223)	88 Radium Ra (226)	**89 Actinium Ac (227)	104 Rutherfordium Rf (261)	105 Dubnium Db (262)	106 Seaborgium Sg (263)	107 Bohrium Bh (261)	108 Hassium Hs (265)	109 Meitnerium Mt (266)	110 *** (269)	111 *** (272)	112 *** (277)						

Inner Transition Elements (f Series)

*Lanthanides 4f	58 Cerium Ce 140.12	59 Praseodymium Pr 140.907	60 Neodymium Nd 144.24	61 Promethium Pm 144.913	62 Samarium Sm 150.35	63 Europium Eu 151.96	64 Gadolinium Gd 157.25	65 Terbium Tb 158.925	66 Dysprosium Dy 162.50	67 Holmium Ho 164.930	68 Erbium Er 167.26	69 Thulium Tm 168.934	70 Ytterbium Yb 173.04	71 Lutetium Lu 174.97
**Actinides 5f	90 Thorium Th 232.038	91 Protactinium Pa (231)	92 Uranium U 238.03	93 Neptunium Np (237)	94 Plutonium Pu 244.064	95 Americium Am (243)	96 Curium Cm (247)	97 Berkelium Bk (247)	98 Californium Cf 242.058	99 Einsteinium Es (254)	100 Fermium Fm 257.095	101 Mendelevium Md 258.10	102 Nobelium No 259.101	103 Lawrencium Lr 260.105

*** These elements have not yet been named.

compounds. A compound is a kind of material formed from two or more elements in which the elements are always combined in the same proportions. Each unit of a particular compound is called a **molecule.** A molecule of a particular compound, for example table sugar ($C_{12}H_{22}O_{11}$), *always* contains 12 atoms of the element carbon, 22 atoms of the element hydrogen, and 11 atoms of the element oxygen. The word *molecule* is used when referring to the numbers of these units, while the word *compound* is used when describing the features or properties of these molecules.

In most cases, elements and compounds are found as **mixtures.** A mixture is matter that contains two or more substances *not* in set proportions. For example, salt water can be composed of varying amounts of NaCl and H_2O. If the components of the mixture are distributed equally throughout it is called a *homogenous solution.* Solutions are homogenous mixtures in which the particles are the size of atoms or small molecules. Another type of mixture called a suspension is similar to a solution. However, the dispersed particles are larger than molecular size. A suspension has particles that eventually separate out and are no longer equally dispersed in the system. Dust particles in the air are an example of a suspension. The dust settles out and collects on tables and other furniture. Another type of mixture is a **colloid.** This system contains dispersed particles that are larger than molecules but still small enough that they do not settle out. Even though colloids are composed of small particles that are mixed together with a liquid such as water, they do not act like solutions or a suspension. In a colloidal system, the dispersed particles form a spongelike network that holds the water molecules in place. One unique characteristic of a colloid is that it can become more or less solid depending on the temperature. When the temperature is lowered, the mixture becomes solidified; as the temperature is increased, it becomes more liquid. We speak of these as the *gel* (solid) and *sol* (liquid) phases of a colloid. A gelatin dessert is a good example of a colloidal system. If you heat the gelatin, it becomes liquid as it changes to the sol phase. If you cool it again, it goes back to the gel phase and becomes solid. Environmental changes other than temperature can also cause colloids to change their phase. In living cells, this sol/gel transformation can cause the cell to move and change shape.

2.2 Structure of the Atom

The smallest part of an element that still acts like that element is called an *atom* and retains all the traits of that element. When we use a **chemical symbol** such as Al for aluminum or C for carbon, it represents one atom of that element. The atom is constructed of three major particles; two of them are in a central region called the **atomic nucleus.** The third type of particle is in the region surrounding the nucleus (figure 2.1). The weight, or mass, of the atom is concentrated in the nucleus. One major group of particles located in the nucleus is the **neutrons;** they were named *neutrons* to reflect their lack of electrical charge. **Protons,** the second type of particle in the nucleus, have a positive electrical charge. **Electrons** fly around the atomic nucleus in certain areas called energy levels and each electron has a negative electrical charge.

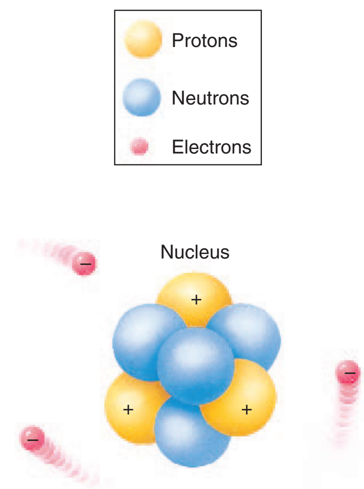


Figure 2.1

Atomic Structure
The nucleus of the atom contains the protons and the neutrons, which are the massive particles of the atom. The electrons, much less massive, are in constant motion about the nucleus. Therefore, the neutrons and protons give an atom its mass (weight) and the volume of an atom is determined by how many and how far out the electrons encircle the neutrons and protons.

An atom is neutral in charge when the number of positively charged protons is balanced by the number of negatively charged electrons. You can determine the number of either of these two particles in a balanced atom if you know the number of the other particle. For instance, hydrogen, with one proton, would have one electron; carbon, with six protons, would have six electrons; and oxygen, with eight electrons, would have eight protons.

The atoms of each kind of element have a specific number of protons. The number of protons determines the identity of the element. For example, carbon always has six protons and no other element has that number. Oxygen always has eight protons. The **atomic number** of an element is the number of protons in an atom of that element; therefore, each element has a unique atomic number. Because oxygen has eight protons, its atomic number is eight. The mass of a proton is 1.67×10^{-24} grams. Because this is an extremely small mass and is awkward to express, it is said to be equal to one **atomic mass unit,** abbreviated **AMU** (table 2.1).

Although all atoms of the same element have the same number of protons, they do not always have the same number of neutrons. In the case of oxygen, over 99% of the atoms have eight neutrons, but there are others with more or

Table 2.1

COMPARISON OF ATOMIC PARTICLES

	Protons	Electrons	Neutrons
Location	Nucleus	Outside nucleus	Nucleus
Charge	Positive (+)	Negative (–)	None (neutral)
Number present	Identical to the atomic number	Equal to number of protons	Atomic weight minus atomic number
Mass	1 AMU	1/1,836 AMU	1 AMU

fewer neutrons. Each atom of the same element with a different number of neutrons is called an **isotope** of an element.

The most common isotope of oxygen has eight neutrons, but another isotope of oxygen has nine neutrons. We can determine the number of neutrons by comparing the masses of the isotopes. The **mass number** or **atomic weight** of an atom is the number of protons plus the number of neutrons in the nucleus. The atomic weights of elements are not whole numbers because the atomic weight is an average of the mass of the different isotopic forms of that element. The atomic weight is customarily used to compare different isotopes of the same element. An oxygen isotope with an atomic weight of 16 AMUs is composed of eight protons and eight neutrons and is identified as ^{16}O. Oxygen 17, or ^{17}O, has a mass of 17 AMUs. Eight of these units are due to the eight protons that every oxygen atom has; the rest of the mass is due to nine neutrons ($17 - 8 = 9$). Figure 2.2 shows different isotopes of hydrogen.

The periodic table of the elements (How Science Works 2.1) lists all the elements in order of increasing atomic number (number of protons). In addition, this table lists the atomic weight of each element. You can use these two numbers to determine the number of the three major particles in an atom—protons, neutrons, and electrons. Look at the periodic table and find helium in the upper-right-hand corner (He). Two is its atomic number; thus, every helium atom will have two protons. Because the protons are positively charged, the nucleus has two positive charges that must be balanced by two negatively charged electrons. The atomic mass of helium is given as 4.0026. This is the calculated average mass of a group of helium atoms. Most of them have a mass of four—two protons and two neutrons. Generally, you will need to work only with the most common isotope, so the atomic weight should be rounded to the nearest whole number. If it is a number like 4.003, use 4 as the most common mass. If the *mass number* is a number like 39.95, use 40 as the nearest whole number. Look at several atoms in the periodic table. You can easily determine the number of protons and the number of neutrons in the most common isotopes of almost all of these atoms.

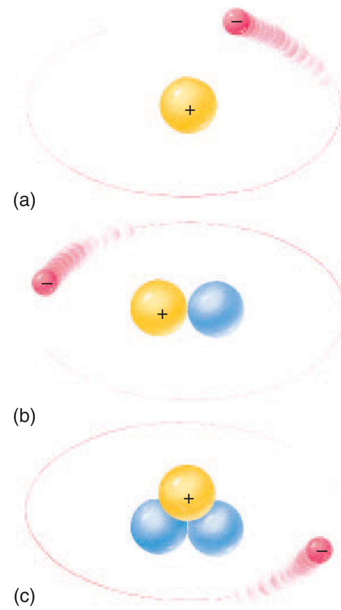


Figure 2.2

Isotopes of Hydrogen
(*a*) The most common form of hydrogen is the isotope that is 1 AMU. It is composed of one proton and no neutrons. (*b*) The isotope *deuterium* is 2 AMU and has one proton and one neutron. (*c*) *Tritium,* 3 AMU, has two neutrons and one proton. Each of these isotopes of hydrogen also has one electron but, because the mass of an electron is so small, the electrons do not contribute significantly to the mass as measured in AMU. All three isotopes of hydrogen are found on Earth, but the most frequently occurring has 1 AMU and is commonly called *hydrogen.* Most scientists use the term "hydrogen" in a generic sense, i.e., the term is not specific but might refer to any or all of these isotopes.

Because isotopes differ in the number of neutrons they contain, the isotopic forms of a particular element differ from one another in some of their characteristics. For example, there are many isotopes of iodine. The most common isotope of iodine is ^{127}I; it has an atomic weight of 127. A different isotope of iodine is ^{131}I; its atomic weight is 131 and it is **radioactive.** This means that it is not stable and that its nucleus disintegrates, releasing energy and particles. The energy can be detected by using photographic film or a Geiger counter. If a physician suspects that a patient has a thyroid gland that is functioning improperly, ^{131}I may be used to help confirm the diagnosis. The thyroid gland, located in a person's neck, normally collects iodine atoms from the blood and uses them in the manufacture of the body-regulating chemical thyroxine. If the thyroid gland is working properly to form thyroxine, the radioactive iodine will collect in the gland, where its presence can be detected.

If no iodine has collected there, the physician knows that the gland is not functioning correctly and can take steps to help the patient.

The number and position of the electrons in an atom are responsible for the way atoms interact with each other. Electrons are the negatively charged particles of an atom that balance the positive charges of the protons in the atomic nucleus. Notice in table 2.1 that the mass of an electron is a tiny fraction of the mass of a proton. This mass is so slight that it usually does not influence the AMU of an element. But electrons are important even though they do not have a major effect on the mass of the element. The number and location of the electrons in any atom determine the kinds of chemical reactions the atom may undergo. All living things have the ability to manipulate matter and energy. In other words, they all have the ability to direct these chemical reactions.

2.3 Chemical Reactions: Compounds and Chemical Change

When atoms or molecules interact with each other and rearrange to form new combinations, we say that they have undergone a **chemical reaction.** A chemical reaction usually involves a change in energy as well as some rearrangement in the molecular structure. We frequently use a chemical shorthand to express what is going on. An arrow (→) indicates that a chemical reaction is occurring. The arrowhead points to the materials that are produced by the reaction; we call these the **products.** On the other side of the arrow, we generally show the materials that are going to react with each other; we call these the ingredients of the reaction or the **reactants.** Some of the most fascinating information we have learned recently concerns the way in which living things manipulate chemical reactions to release or store chemical energy. This material is covered in detail in chapters 5 and 6.

Figure 2.3 shows the chemical shorthand used to indicate several reactions. The chemical shorthand is called an *equation.* Look closely at the equations and identify the reactants and products in each. Six of the most important chemical reactions that occur in organisms are (1) hydrolysis (breaking a molecule using a water molecule), (2) dehydration synthesis (combining smaller molecules by extracting the equivalent of water molecules from the parts), (3) oxidation-reduction (reactions that may release or store energy), (4) acid-base (reaction between an acid and a base), (5) phosphorylation (adding a phosphate), and (6) transfer (switching partners).

Electron Distribution

Electrons are constantly moving at great speeds and tend to be found in specific regions some distance from the nucleus (figure 2.4). The position of an electron at any instant in time is determined by several factors. First, because protons

$C_6H_{12}O_6 + 6\ O_2 \rightarrow 6\ CO_2 + 6\ H_2O + \text{energy}$	Oxidation-reduction
$HCl + NaOH \rightarrow NaCl + H_2O$	Acid-base (neutralization)
$C_6H_{12}O_6 + C_6H_{12}O_6 \rightarrow C_{12}H_{22}O_{11} + H_2O$	Dehydration synthesis
$C_{12}H_{22}O_{11} + H_2O \rightarrow 2\ C_6H_{12}O_6$	Hydrolysis
$C_6H_{12}O_6 + PO_4^{3-} \rightarrow C_6H_{11}O_6PO_4^{2-} + H^+$	Phosphorylation
$Cd(NO_3)_2 + Na_2S \rightarrow CdS + 2NaNO_3$	Transfer

Figure 2.3

Chemical Equations
The equations here use chemical shorthand to indicate that there has been a rearrangement of the chemical bonds in the reactants to form the products. Along with the rearrangement of the chemical bonds, there has been a change in the energy content. Notice the numbers in front of the formula for oxygen, carbon dioxide, and water (e.g., 6 H_2O). That number indicates that there are a total of six water molecules formed in this reaction. If there is no such number preceding a formula, it is assumed that the number is one (1) of that kind of unit.

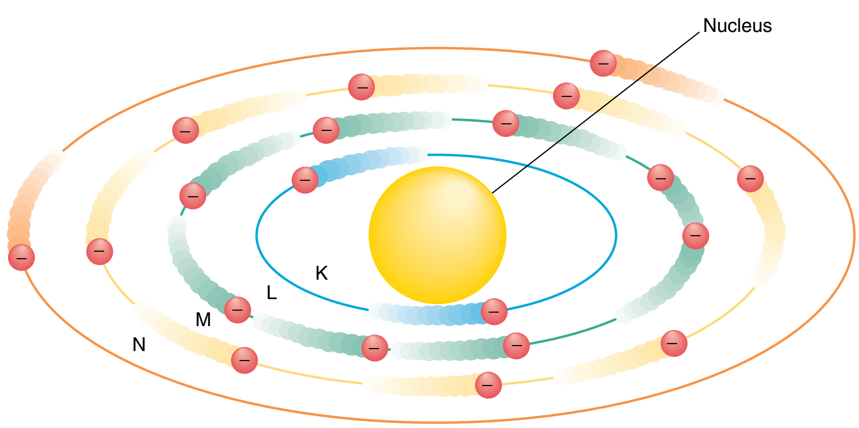
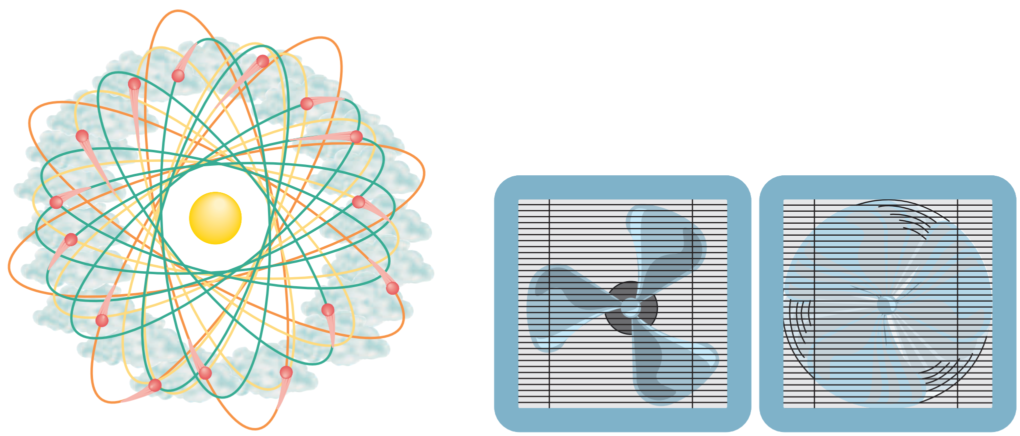


Figure 2.4

The Bohr Atom
Several decades ago it was thought that electrons revolved around the nucleus of the atom in particular paths, or tracks. Each track was labeled with a letter: K, L, M, N, and so on. Each track was thought to be able to hold a specific number of electrons moving at a particular speed. These electron tracks were described as quanta of energy.



(a) (b)

Figure 2.5

The Electron Cloud
So fast are the electrons moving around the nucleus that they can be thought of as forming a cloud around it, rather than an orbit or single track. (*a*) You might think of the electron cloud as hundreds of photographs of an atom. Each photograph shows where an electron was at the time the picture was taken. But when the next picture is taken, the electron is somewhere else. In effect, an electron is everyplace in its energy level at the same time just as the fan blade of a window fan is everywhere at once when it is running (*b*). No matter where you stick your finger, you will be touched (ouch!) by the moving blade should you stick your finger in the fan! Although we are able to determine where an electron is at a given time, we do not know the path it uses to go from one place to another.

and electrons are of opposite charge, electrons are attracted to the protons in the nucleus of the atom. Second, counterbalancing this is the force created by the movement of the electrons, which tends to cause them to move away from the nucleus. Third, the electrons repel one another because they have identical negative charges. The balance of these three forces creates a situation in which the electrons of an atom tend to remain in the neighborhood of the nucleus but are kept apart from one another. Electron distribution is not random; electrons are likely to be found in certain locations.

When chemists first described the atom, they tried to account for the fact that electrons seemed to be traveling at one of several different speeds about the atomic nucleus. Electrons did not travel at intermediate speeds. Because of this, it was thought that electrons followed a particular path, or orbit, similar to the orbits of the planets about the Sun.

A Model of the Atom

Several decades ago, as more experimental data were gathered and interpreted, we began to formulate a model for the structure of atoms. In this model, each region, called an **energy level,** contains electrons moving at approximately the same speed. These electrons also have about the same amount of kinetic energy. Each energy level is numbered in increasing order, that is, energy level 1 contains electrons with the lowest amount of energy, energy level 2 has electrons with more energy than those found in energy level 1, energy level 3 has electrons with even more energy than those in level 2, and so forth. It was also found that electrons do not encircle the atomic nucleus in flat, two-dimensional paths. Some move around the atomic nucleus in a three-dimensional region that is spherical, forming cloudlike layers about the nucleus (figure 2.5). Others move in a manner that resembles the figure eight (8), forming cloudlike regions that look like dumbbells or hourglasses. No matter how many electrons in an energy level or what shape path they follow, all the electrons in a single energy level contain approximately the same amount of kinetic energy.

For most biologically important atoms, the number of electrons in the first energy level can contain two electrons, the second energy level can contain a total of eight electrons, the third energy level eight, and so forth (table 2.2).

Notice in table 2.2 that the number of protons in each atomic nucleus equals the total number of electrons moving about it. Also note that some of the elements (unshaded areas) are atoms with outermost energy levels that contain the maximum number of electrons they can hold, for example, He, Ne, Ar. Elements such as He and Ne with filled outer energy levels are particularly stable. Atoms have a tendency to seek such a stable, filled outer energy level arrangement, a tendency referred to as the *octet (8) rule*. The rule states that atoms attempt to acquire an outermost energy level with eight electrons through chemical reactions. Since elements like He and Ne have full outermost energy levels under ordinary circumstances, they do not normally undergo

Table 2.2

THE NUMBER OF ELECTRONS POSSIBLE IN ENERGY LEVELS

Element	Symbol	Atomic Number	Number of Electrons Required to Fill Each Energy Level			
			Energy Level 1	Energy Level 2	Energy Level 3	Energy Level 4
Hydrogen	H	1	1			
Helium	He	2	2			
Carbon	C	6	2	4		
Nitrogen	N	7	2	5		
Oxygen	O	8	2	6		
Neon	Ne	10	2	8		
Sodium	Na	11	2	8	1	
Magnesium	Mg	12	2	8	2	
Phosphorus	P	15	2	8	5	
Sulfur	S	16	2	8	6	
Argon	Ar	18	2	8	8	
Chlorine	Cl	17	2	8	7	
Potassium	K	19	2	8	8	1
Calcium	Ca	20	2	8	8	2

chemical reactions and are therefore referred to as noble or inert. Atoms of other elements have outer energy levels that are not full, for example, H, C, Mg, and will undergo reactions to fill their outermost energy level in order to become stable.

Ions

Remember that atoms are electrically neutral when they have equal numbers of protons and electrons. Certain atoms, however, are able to exist with an unbalanced charge; that is, the number of protons is not equal to the number of electrons. These unbalanced, or charged, atoms are called **ions.**

The ion of sodium, for example, is formed when 1 of the 11 electrons of the sodium atom escapes. It tends to lose this electron in order to become more stable, that is, follow the octet rule. The sodium nucleus is composed of 11 positive charges (protons) and 12 neutrons. (The most common isotope of sodium is sodium 23, which has 12 neutrons.) The 11 electrons that balance the charge are most likely positioned as follows: 2 electrons in the first energy level, 8 in the second energy level, and 1 in the third energy level. Focus your attention on the outermost electron. For an atom of sodium to follow the octet rule it has two choices: it can either (1) gain 7 new electrons to fill the third energy level or (2) lose this single outermost electron, thus making the second energy level the outermost and full with eight electrons. Sodium typically loses this last third energy electron to fulfill the octet rule (figure 2.6*A*). What remains when the electron leaves the atom is called the *ion.* In this case, the sodium ion is now composed of the 11 positively charged protons and the 12 neutral neutrons—but it has only 10 electrons. The fact that there are 11 positive and only 10 negative charges means that there is an excess of 1 positive charge. This sodium ion now has its outermost energy level full of electrons, that is, it contains eight electrons. In this state, the atom is electrically charged, but more stable. All positively charged ions are called **cations.** We still use the chemical symbol Na to represent the ion, but we add the superscript^{+} to indicate that it is no longer a neutral atom but an electrically charged ion (Na^{+}). It is easy to remember that a cation (positive ion) is formed because it *loses* negative electrons.

Some atoms become more stable by *acquiring* one or more electrons in their outermost energy levels. For example, the outermost energy level of an atom of oxygen contains six electrons. It would be more stable if it had eight. In this case, an atom of oxygen may acquire these two electrons from another atom that would serve as an electron donor. When these two electrons are acquired, an atom of oxygen becomes an ion of oxygen and has a double negative charge ($O^{=}$). Negatively charged ions are referred to as **anions.**

The sodium ion is relatively stable because its outermost energy level is full. A sodium atom will lose one electron from its third major energy level so that the second energy level becomes outermost and is full of electrons. Similarly, magnesium loses two electrons from its third major energy level so that the second major energy level, which is full with eight electrons, becomes outermost. When a magnesium atom (Mg) loses two electrons, it becomes a magnesium ion (Mg^{++}). The periodic table of the elements is arranged so that all atoms in the first column become ions in a similar way. That is, when they form ions, they do so by losing one electron. Each becomes a $^{+}$ ion. Atoms in the second column of the periodic table become $^{++}$ ions when they lose two electrons. Atoms at the extreme right of the periodic table of the

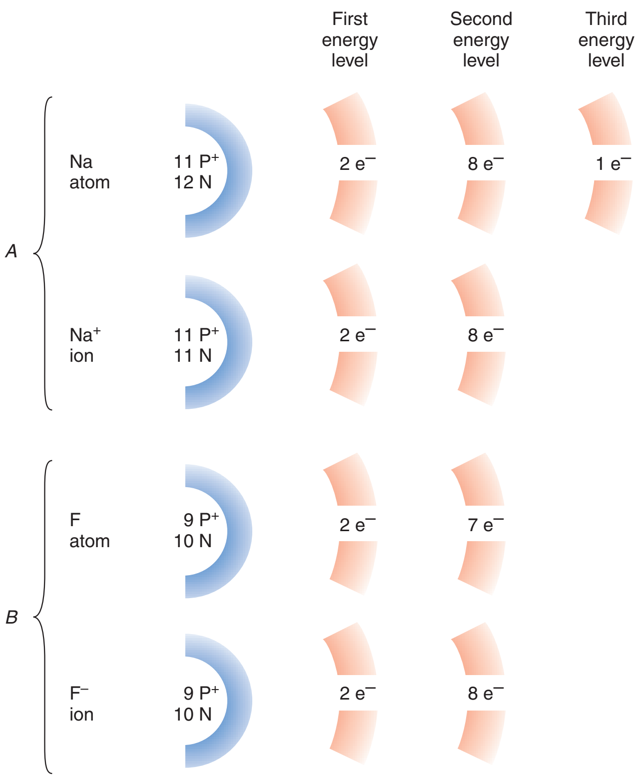


Figure 2.6

Ion Formation

A sodium atom (*A*) has two electrons in the first energy level, eight in the second energy level, and one in the third level. When it loses its one outer electron, it becomes a sodium ion. An atom of fluorine (*B*) has two electrons in the first energy level and only seven in the second energy level. To become stable, fluorine picks up an extra electron from an electron donor to fill its outermost energy level, thus satisfying the octet rule.

elements do not become ions; they tend to be stable as atoms. These atoms are called *inert* or *noble* because of their lack of activity. They seldom react because their protons and electrons are equal in number and they have a full outer energy level; therefore, they are not likely to lose electrons (table 2.2).

The column to the left of these gases contains atoms that lack a full outer energy level. They all require an additional electron. Fluorine with its nine electrons would have two in the first energy level and seven electrons in the second energy level. The second major energy level can hold a total of eight electrons. You can see that one additional electron could fit into the second energy level. Whenever the atom of fluorine can, it will accept an extra electron so that its outermost energy level is full. When it does so, it no longer has a balanced charge. When it accepts an extra electron, it has one more negative electron than positive protons; thus, it has become a negative ion (F^-) (figure 2.6*b*).

Similarly, chlorine will form a $^-$ ion, anion. Oxygen, in the next column, will accept two electrons and become a negative ion with two extra negative charges ($O^=$). If you know the number and position of the electrons, you are better able to hypothesize whether or not an atom will become an ion and, if it does, whether it will be a positive ion or a negative ion. You can use the periodic table of the elements to help you determine an atom's ability to form ions. This information is useful as we see how ions react to each other.

2.4 Chemical Bonds

There are a variety of physical and chemical forces that act on atoms and make them attractive to each other. Each of these results in a particular arrangement of atoms or association of atoms. The forces that combine atoms and hold them together are called **chemical bonds.** Bonds are formed in an attempt to stabilize atoms energetically, that is, complete their outer shells. There are two major types of chemical bonds. They differ from one another with respect to the kinds of attractive forces holding the atoms together. The bonding together of atoms results in the formation of a molecule of a compound. This molecule is composed of a specific number of atoms (or ions) joined to each other in a particular way and is represented by a **chemical formula.** We generally use the chemical symbols for each of the component atoms when we designate a molecule. Sometimes there will be a small number following the chemical symbol. This number indicates how many atoms of that particular element are used in the molecule. The group of chemical symbols and numbers is termed an **empirical formula;** it will tell you what elements are in a compound and also how many atoms of each element are required. For example, $CaCl_2$ tells us that a molecule of calcium chloride is composed of one calcium atom and two chlorine atoms. A **structural formula** is a drawing that shows not only the kinds of atoms in the molecule but also the number and spacial arrangement of atoms within the molecule.

The properties of a compound are very different from the properties of the atoms that make up the compound. Table salt is composed of the elements sodium (a silvery-white, soft metal) and chlorine (a yellowish-green gas) bound together. Both sodium and chlorine are very dangerous when they are by themselves. When they are combined as salt, the compound is a nontoxic substance, essential for living organisms.

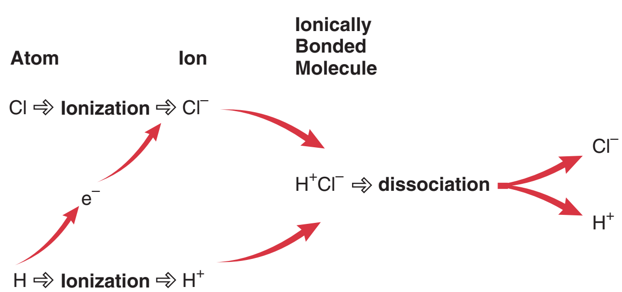
Ionic Bonds

When positive and negative ions are near each other, they are mutually attracted because of their opposite charges. This attraction between ions of opposite charge results in the formation of a stable group of ions. This force of attraction is termed an **ionic bond.** Compounds that form as a result of attractions between ions are called *ionic compounds* and are very important in living systems. We can categorize these ionic compounds into three different groups.

$$Ca + \begin{matrix} Cl \\ Cl \end{matrix} \rightarrow Ca^{2+} \begin{matrix} (Cl)^- \\ (Cl)^- \end{matrix}$$

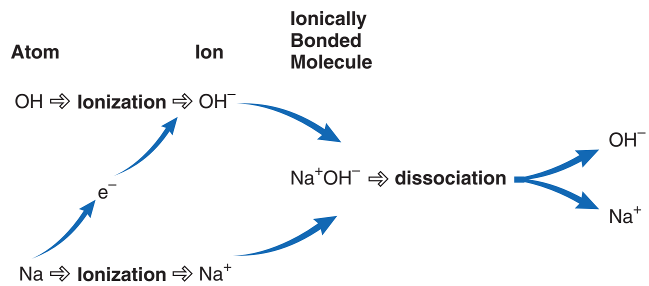
Acids, Bases, and Salts

Acids and bases are two classes of biologically important compounds. Their characteristics are determined by the nature of their chemical bonds. When acids are mixed in water, hydrogen ions (H^+) are set free. The hydrogen ion is positive because it has lost its electron and now has only the positive charge of the proton. An **acid** is any ionic compound that releases a hydrogen ion in a solution. You can think of an acid, then, as a substance able to donate a proton to a solution. However, this is only part of the definition of an acid. We also think of acids as compounds that act like the hydrogen ion—they attract negatively charged particles. Acids have a sour taste such as the taste of citrus fruits. However, tasting chemicals to see if they are acids can be very hazardous since many are highly corrosive. An example of a common acid with which you are probably familiar is the sulfuric acid—$(H^+)_2(SO_4^=)$—in your automobile battery.



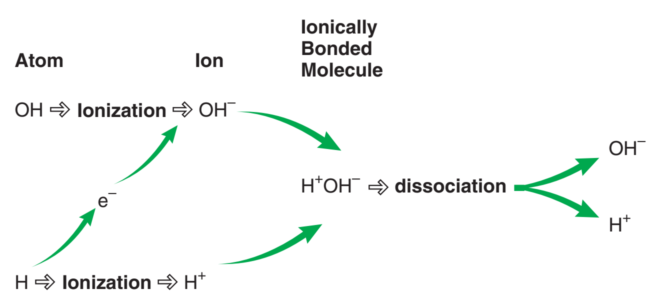
Acids are ionically bonded molecules which when placed in water dissociate into hydrogen (H^+) ions.

Bases or alkaline substances have a slippery feel on the skin. They have a caustic action on living tissue, changing it into a soluble substance. A strong base is used to react with fat to make soap, giving soap its slippery feeling. Bases are also used in certain kinds of batteries, that is, alkaline batteries. Weak bases have a bitter taste, for example, the taste of coffee. A **base** is the opposite of an acid in that it is an ionic compound that releases a group known as a **hydroxide ion,** or OH^- group. This group is composed of an oxygen atom and a hydrogen atom bonded together, but with an additional electron. The hydroxide ion is negatively charged. It is a base because it is able to donate electrons to the solution. A base can also be thought of as any substance that is able to attract positively charged hydrogen ions (H^+). A very strong base used in oven cleaners is $Na^+(OH)^-$, sodium hydroxide. Notice that free ions are always written with the type and number of their electrical charge as a superscript.



Basic (alkaline) substances are ionically bonded molecules which when placed in water dissociate into hydroxide (OH^-) ions.

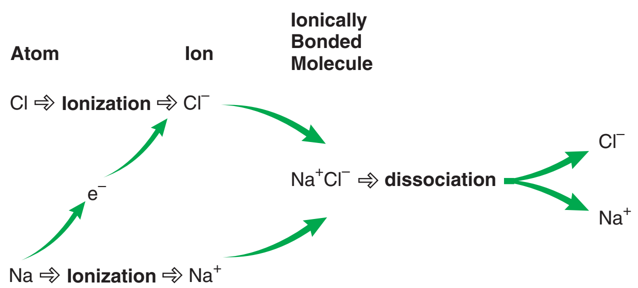
The degree to which a solution is acidic or basic is represented by a quantity known as **pH.** The pH scale is a measure of hydrogen ion concentration. A pH of 7 indicates that the solution is neutral and has an equal number of H^+ ions and OH^- ions to balance each other. As the pH number gets smaller, the number of hydrogen ions in the solution increases. A number higher than 7 indicates that the solution has more OH^- than H^+. As the pH number gets larger, the number of hydroxide ions increases (figure 2.7).



When water dissociates it releases both hydrogen (H^+) and hydroxide (OH^-) ions. It is neither a base nor an acid. Its pH is 7, neutral.

An additional group of biologically important ionic compounds is called the *salts.* **Salts** are compounds that do not release either H^+ or OH^-; thus, they are neither acids nor bases. They are generally the result of the reaction between an acid and a base in a solution. For example, when an acid such as HCl is mixed with NaOH in water, the H^+ and the OH^- combine with each other to form water, H_2O. The remaining ions (Na^+ and Cl^-) join to form the salt NaCl:

$$\mathbf{HCl + NaOH \rightarrow (Na^+ + Cl^- + H^+ + OH^-) \rightarrow NaCl + H_2O}$$



Salts are ionically bonded molecules which when placed in water dissociate into ions that are neither hydrogen (H^+) or hydroxide (OH^-).

The chemical reaction that occurs when acids and bases react with each other is called **neutralization.** The acid no longer acts as an acid (it has been neutralized) and the base no longer acts as a base.

As you can see from figure 2.7, not all acids or bases produce the same pH. Some compounds release hydrogen ions very easily, cause low pHs, and are called *strong acids*. Hydrochloric acid (HCl) and sulfuric acid (H_2SO_4) are examples of strong acids. Many other compounds give up their hydrogen ions grudgingly and therefore do not change the pH very much. They are known as *weak acids*. Carbonic acid (H_2CO_3) and many organic acids found in living things are weak acids. Similarly, there are strong bases like sodium hydroxide (NaOH) and weak bases like sodium bicarbonate—$Na^+(HCO_3)^-$.

Covalent Bonds

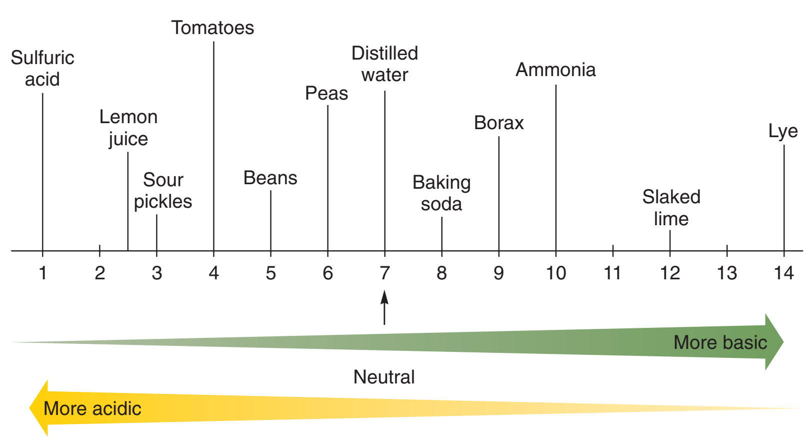
In addition to ionic bonds, there is a second strong chemical bond known as a *covalent bond*. A **covalent bond** is formed when two atoms share a pair of electrons. This sharing can occur when the outermost energy levels of two atoms come close enough to allow the electrons of one to fly around the outermost energy level of the other. These two atoms have energy levels that overlap one another. A covalent bond should be thought of as belonging to each of the atoms involved. You can visualize the bond as people shaking hands: the people are the atoms, the hands are electrons to be shared, and the handshake is the combining force (figure 2.8). Generally, this sharing of a pair of electrons is represented by a single straight line between the atoms involved. The reason covalent bonds form relates to the arrangement of electrons within the atoms. There are many elements that do not tend to form ions. They will not lose electrons, nor will they gain electrons. Instead, these elements get close enough to other atoms that have unfilled energy levels and share electrons with them. If the two elements have orbitals that overlap, the electrons can be shared. By sharing electrons, each atom fills its unfilled outer energy level. Both atoms become more stable as a result of the formation of this covalent bond.

Molecules are defined as the smallest particles of chemical compounds. They are composed of a specific number of atoms arranged in a particular pattern. For example, a molecule of water is composed of one oxygen atom bonded covalently to two atoms of hydrogen. The shared electrons are in the second energy level of oxygen, and the bonds are almost at right angles to each other. Now that you realize how and why bonds are formed, it makes sense that only certain numbers of certain atoms will bond with one another to form molecules. Chemists also use the term *molecule* to mean the smallest naturally occurring part of an element or compound. Using this definition, one atom of iron is a molecule because one atom is the smallest natural piece of the element. Hydrogen, nitrogen, and oxygen tend to form into groups of two atoms. Molecules of these elements are composed of two atoms of hydrogen, two atoms of nitrogen, and two atoms of oxygen, respectively.

H:H	N⋮N	O⁚O
H-H	N≡N	O=O
H_2	N_2	O_2

Figure 2.7

The pH Scale
The concentration of acid (proton donor or electron acceptor) is greatest when the pH number is lowest. As the pH number increases, the concentration of base (proton acceptor or electron donor) increases. At a pH of 7.0, the concentrations of H^+ and OH^- are equal. We usually say, as the pH number gets smaller the solution becomes more acid. As the pH number gets larger the solution becomes more basic or alkaline.



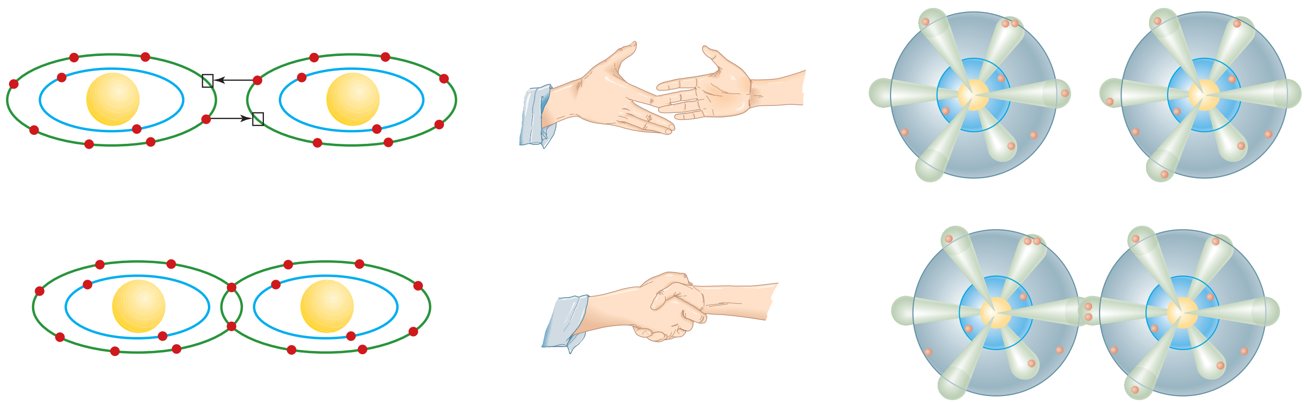


Figure 2.8

Covalent Bonds
When two atoms come sufficiently close to each other that the locations of the outermost electrons overlap, an electron from each one can be shared to "fill" that outermost energy-level area. When two people shake hands, they need to be close enough to each other so that their hands can overlap. At the left, using the Bohr model, the *L*-shells of the two atoms overlap, and so each shell appears to be full. Using the modern model at the right, the propeller-shaped orbitals of the second energy level of each atom overlap, so that each energy level is full. Notice that just as it takes two hands to form a handclasp, it takes two electrons to form a covalent bond.

Hydrogen Bonds

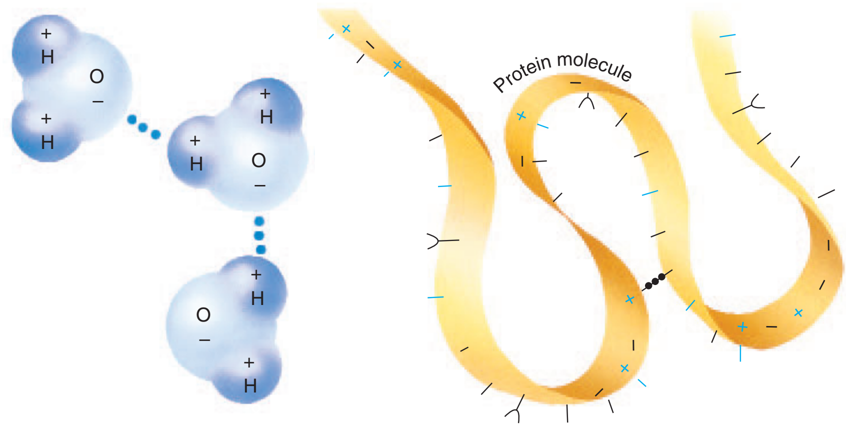
Molecules that are composed of several atoms sometimes have an uneven distribution of charge. This may occur because the electrons involved in the formation of bonds may be located on one side of the molecule. This makes that side of the molecule slightly negative and the other side slightly positive. One side of the molecule has possession of the electrons more than the other side. When a molecule is composed of several atoms that have this uneven charge distribution, the whole molecule may show a positive side and a negative side. We sometimes think of such a molecule as a tiny magnet with a positive pole and a negative pole. This polarity of the molecule may influence how the molecule reacts with other molecules. When several of these **polar molecules** are together, they orient themselves so that the slightly positive end of one is near the slightly negative end of another.

This intermolecular (i.e., between molecules) force of attraction is referred to as a **hydrogen bond.** However, the term *bond* in its purest sense refers only to ionic and covalent forces which hold atoms together to form molecules. Hydrogen bonds hold molecules together; they do not bond atoms together. Because hydrogen has the least attractive force for electrons when it is combined with other elements, the hydrogen electron tends to spend more of its time encircling the other atom's nucleus than its own. The result is the formation of a polar molecule. When the negative pole of this molecule is attracted to the positive pole of another similar polar molecule, the hydrogen will usually be located between the two molecules. Because the hydrogen serves as a bridge between the two molecules, this weak bond has become known as a hydrogen bond.

We usually represent this attraction as three dots between the attracted regions. This weak bond is not responsible for forming molecules, but it is important in determining how groups of molecules are arranged. Water, for example, is composed of polar molecules that form hydrogen bonds (figure 2.9 left). Because of this, individual water molecules are less likely to separate from each other. They need a large input of energy to become separated. This is reflected in the relatively high boiling point of water in comparison to other substances, such as rubbing alcohol. In addition, when a very large molecule, such as a protein or DNA (which is long and threadlike), has parts of its structure slightly positive and other parts slightly negative, these two areas will attract each other and result in coiling or folding of the molecule in particular ways (figure 2.9 right).

Figure 2.9

Hydrogen Bonds
Water molecules arrange themselves so that their positive portions are near the negative portions of other water molecules. The attraction force of a single hydrogen bond is indicated as three dots in a row. It is this kind of intermolecular bonding that accounts for water's unique chemical and physical properties. Without such bonds, life as we know it on Earth would be impossible. The large protein molecule here also has polar areas. When the molecule is folded so that the partially positive areas are near the partially negative areas, a slight attraction forms that tends to keep it folded.



SUMMARY

All matter is composed of atoms, which contain a nucleus of neutrons and protons. The nucleus is surrounded by moving electrons. There are many kinds of atoms, called elements. These differ from one another by the number of protons and electrons they contain. Each is given an atomic number, based on the number of protons in the nucleus, and an atomic weight, determined by the total number of protons and neutrons. Atoms of an element that have the same atomic number but differ in their atomic weight are called isotopes. Some isotopes are radioactive, which means that they fall apart, releasing energy and smaller, more stable particles. Atoms may be combined into larger units called molecules. Two kinds of chemical bonds allow molecules to form—ionic bonds and covalent bonds. A third bond, the hydrogen bond, is a weaker bond that holds molecules together and may also help large molecules maintain a specific shape.

Energy can neither be created nor destroyed, but it can be converted from one form to another. Potential energy and kinetic energy can be interconverted. When energy is converted from one form to another, some of the useful energy is lost. The amount of kinetic energy that the molecules of various substances contain determines whether they are solids, liquids, or gases. The random motion of molecules, which is due to their kinetic energy, results in their distribution throughout available space.

An ion is an atom that is electrically unbalanced. Ions interact to form ionic compounds, such as acids, bases, and salts. Compounds that release hydrogen ions when mixed in water are called acids; those that release hydroxide ions are called bases. A measure of the hydrogen ions present in a solution is known as the pH of the solution. Molecules that interact and exchange parts are said to undergo chemical reactions. The changing of chemical bonds in a reaction may release energy or require the input of additional energy.

THINKING CRITICALLY

Sodium bicarbonate ($NaHCO_3$) is a common household chemical known as baking soda, bicarbonate of soda, or bicarb. It has many uses other than baking. It is a component of many products including toothpaste and antacids, swimming pool chemicals, and headache remedies. When baking soda comes in contact with hydrochloric acid, the following reaction occurs:

$$HCl + NaHCO_3 \rightarrow NaCl + CO_2 + H_2O$$

Can you describe what happens to the atoms in this reaction? In your description, include changes in chemical bonds, pH, and kinetic energy. Can you describe why the baking soda is such an effective chemical in the above-mentioned products? You might try this at home: place a pinch of sodium bicarbonate ($NaHCO_3$) on a plate. Add a couple of drops of vinegar. Observe the reaction. Based on the reaction above, can you explain chemically what has happened?

CONCEPT MAP TERMINOLOGY

Construct a concept map to show relationships among the following concepts.

anion
cation
electron
ion
ionic bond
molecule
neutron
proton
salt

KEY TERMS

acid
anion
atom
atomic mass unit (AMU)
atomic nucleus
atomic number
atomic weight (mass number)
base
cation
chemical bonds
chemical formula
chemical reaction
chemical symbol
colloid
compound
covalent bond
density
electrons
elements
empirical formula
energy level
first law of thermodynamics

gas
hydrogen bond
hydroxide ion
ionic bond
ions
isotopes
kinetic energy
liquid
mass number
matter
mixture
molecule
neutralization
neutrons
periodic table of the elements
pH
polar molecule
potential energy
products
protons
radioactive
reactants
salts
solid
states of matter
structural formula
temperature

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Topics	Questions	Media Resources
2.1 The Basics: Matter and Energy	1. What is the difference between an atom and an element? 2. What is the difference between a molecule and a compound?	**Quick Overview** • Chemistry basics **Key Points** • The basics: Matter and energy **Animations and Review** • Basic chemistry **Interactive Concept Maps** • Ways of looking at matter
2.2 Structure of the Atom	3. How many protons, electrons, and neutrons are in a neutral atom of potassium having an atomic weight of 39? 4. Diagram an atom showing the positions of electrons, protons, and neutrons. 5. Diagram two isotopes of oxygen. 6. Define the following terms: AMU and atomic number.	**Quick Overview** • The parts of an atom **Key Points** • Structure of the atom **Animations and Review** • Atoms **Interactive Concept Maps** • Information from the periodic table • Subatomic particles
2.3 Chemical Reactions: Compounds and Chemical Change	7. Define the term: second energy level. 8. What is the difference between a cation and an anion?	**Quick Overview** • Elements and compounds **Key Points** • Chemical reactions: Compounds and chemical change
2.4 Chemical Bonds	9. Define the terms: polar molecule and covalent bond. 10. Name three kinds of chemical bonds that hold atoms or molecules together. How do these bonds differ from one another? 11. What does it mean if a solution has a pH number of 3, 12, 2, 7, or 9? 12. What relationship does kinetic energy have to the three states of matter? homogenous solutions? chemical bonds? 13. Define the term: chemical reaction, and give an example.	**Quick Overview** • Different types of chemical bonds **Key Points** • Chemical bonds **Animations and Review** • Bonds • Water • pH **Interactive Concept Maps** • Text concept map **Experience This!** • Hydrogen bonds and surface tension