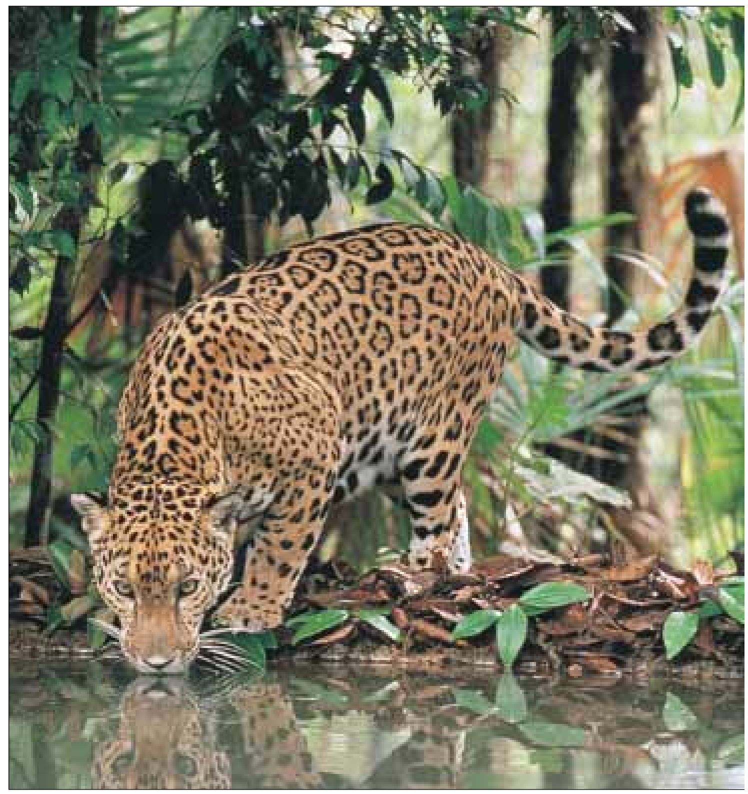
CHAPTER 2

Atoms and Molecules: The Chemical Basis of Life



(Frans Lanting/Minden Pictures)

This jaguar and the plants of the rain forest, as well as an abundance of insects and microorganisms, share fundamental similarities in their chemical composition and basic metabolic processes. These chemical similarities provide strong evidence for the evolution of all organisms from a common ancestor and explain why much of what biologists learn from studying bacteria or rats in laboratories can be applied to other organisms, including humans. Furthermore, the basic physical principles governing organisms are not unique to living things, for they apply to nonliving systems as well.

Today much attention is given to **molecular biology**—the chemistry and physics of the molecules that constitute living things. A molecular biologist might study how proteins interact with DNA in ways that control the expression of certain genes, or might investigate the precise interactions among a cell's atoms and molecules that maintain the energy flow essential to life. However, an understanding of chemistry is essential to *all* biologists. An evolutionary biologist might study evolutionary relationships by comparing proteins produced by different types of organisms. An ecologist might study the biological effects of changes in the salinity of the water in an estuary. A botanist might be a "chemical prospector," seeking new sources of medicines from plants.

In this chapter we lay a foundation for understanding how the structure of atoms determines the way they form chemical bonds to produce complex compounds. Most of our discussion will center around small, simple substances known as **inorganic compounds.** Among the biologically important groups of inorganic compounds are water, simple acids and bases, and simple salts. We pay particular attention to water, the most abundant substance on Earth's surface and in organisms, and we examine how its unique properties affect living things as well as their nonliving environment. In Chapter 3 we extend our discussion to **organic compounds,** which are generally large and complex and which always contain carbon atoms joined together to form the backbone, or skeleton, of the molecule.

LEARNING OBJECTIVES

AFTER YOU HAVE STUDIED THIS CHAPTER YOU SHOULD BE ABLE TO

1. Name the principal chemical elements in living things and give an important function of each.
2. Compare the physical properties (mass and charge) and the locations of electrons, protons, and neutrons.
3. Distinguish between the atomic number and the mass number of an element.
4. Define the term *electron orbital,* and relate orbitals to energy levels.
5. Explain how the number of valence electrons of an atom is related to its chemical properties.
6. Distinguish among covalent bonds, hydrogen bonds, and ionic bonds. Compare them in terms of the mechanisms by which they form and their relative bond strengths.
7. Explain how cations and anions form and how they interact.
8. Distinguish between the terms *oxidation* and *reduction* and relate these processes to the transfer of energy.
9. Draw a simple ball-and-stick model of a water molecule, indicating the regions of partial positive and partial negative charge. Show how hydrogen bonds form between adjacent water molecules and explain how these are responsible for many of the properties of water.
10. Contrast acids and bases and discuss their properties.
11. Convert the hydrogen ion concentration (moles per liter) of a solution to a pH value. Describe how buffers help minimize changes in pH.
12. Describe the composition of a salt and explain why salts are important in organisms.

ELEMENTS ARE NOT CHANGED IN NORMAL CHEMICAL REACTIONS

Elements are substances that cannot be broken down into simpler substances by ordinary chemical reactions. Scientists have assigned each element a **chemical symbol:** usually the first letter or first and second letters of the English or Latin name of the element. For example, O is the symbol for oxygen, C for carbon, H for hydrogen, N for nitrogen, and Na for sodium (Latin *natrium*).

Just four elements—oxygen, carbon, hydrogen, and nitrogen—are responsible for over 96% of the mass of most organisms. Others, such as calcium, phosphorus, potassium, and magnesium, are also consistently present but in smaller quantities. Some elements, such as iodine and copper, are known as *trace elements* because they are present only in minute amounts. Table 2–1 lists the elements that make up two representative organisms, a human and a typical nonwoody plant, and briefly explains why each is important.

ATOMS ARE THE FUNDAMENTAL PARTICLES OF ELEMENTS

An **atom** is the smallest portion of an element that retains its chemical properties. Atoms are much smaller than the tiniest particle visible under a light microscope. By special scanning tunneling electron microscopy, with magnification as high as × 5 million, researchers have been able to photograph the positions of some large atoms in molecules.

Physicists have discovered a number of subatomic particles, but for our purposes we need consider only three: electrons, protons, and neutrons. An **electron** is a particle that carries a unit of negative electrical charge; a **proton** carries a unit of positive charge; and a **neutron** is an uncharged particle. In an electrically neutral atom, the number of electrons is equal to the number of protons.

Clustered together, protons and neutrons compose the **atomic nucleus.** Electrons, however, have no fixed locations and move rapidly through space outside the atomic nucleus.

An atom is uniquely identified by its number of protons

Each kind of element has a fixed number of protons in the atomic nucleus. This number, called the **atomic number,** is written as a subscript to the left of the chemical symbol. Thus $_{1}H$ indicates that the hydrogen nucleus contains one proton, and $_{8}O$ that the oxygen nucleus contains eight protons. It is the atomic number, the number of protons in its nucleus, that determines an atom's identity.

The **periodic table** (Fig. 2–1 and Appendix B) is a chart in which elements are arranged in order by atomic number. As will become evident in subsequent discussions, the periodic table is an extremely useful device because it allows us to simultaneously correlate a great many of the relationships among the various elements.

Figure 2–1 includes representations of the **electron configurations** of several elements important in organisms. These *Bohr models,* which show the electrons arranged in a series of concentric circles around the nucleus, are convenient to use, but inaccurate. As we will see, electrons do not actually circle the nucleus in fixed concentric pathways.

Protons plus neutrons determine atomic mass

The mass of a subatomic particle is exceedingly small, much too small to be conveniently expressed in grams or even micrograms. Such masses are expressed in terms of the **atomic**

Table 2–1 Elements That Make Up Some Representative Organisms

Element and Chemical Symbol	Approximate % of Total Mass of Human Body	Approximate % of Total Mass of Nonwoody Plant	Importance or Functions
Oxygen (O)	65	78	Required for cellular respiration; present in most organic compounds; component of water
Carbon (C)	18	11	Forms backbone of organic molecules; each carbon atom can form four bonds with other atoms
Hydrogen (H)	10	9	Present in most organic compounds; component of water; hydrogen ion (H^+) is involved in some energy transfers
Nitrogen (N)	3	*	Component of proteins and nucleic acids; component of chlorophyll in plants
Calcium (Ca)	1.5	*	Structural component of bones and teeth; calcium ion (Ca^{2+}) is important in muscle contraction, conduction of nerve impulses, and blood clotting; associated with plant cell wall
Phosphorus (P)	1	*	Component of nucleic acids and of phospholipids in membranes; important in energy transfer reactions; structural component of bone
Potassium (K)	*	*	Potassium ion (K^+) is principal positive ion (cation) in interstitial (tissue) fluid of animals; important in nerve function; affects muscle contraction; controls opening of stomata in plants
Sulfur (S)	*	*	Component of most proteins
Sodium (Na)	*	*	Sodium ion (Na^+) is principal positive ion (cation) in interstitial (tissue) fluid of animals; important in fluid balance; essential for conduction of nerve impulses; not essential in most plants
Magnesium (Mg)	*	*	Needed in blood and other tissues of animals; activates many enzymes; component of chlorophyll in plants
Chlorine (Cl)	*	*	Chloride ion (Cl^-) is principal negative ion (anion) in interstitial (tissue) fluid of animals; important in water balance; essential for photosynthesis
Iron (Fe)	*	*	Component of hemoglobin in animals; activates certain enzymes

*The asterisk indicates that these elements represent less than 1% of the total mass. Other elements found in very small (trace) amounts in animals, plants, or both include iodine (I), manganese (Mn), copper (Cu), zinc (Zn), cobalt (Co), fluorine (F), molybdenum (Mo), selenium (Se), boron (B), silicon (Si), and a few others.

mass unit (amu), also called the **dalton** in honor of John Dalton, who formulated an atomic theory in the early 1800s. One amu is equal to the approximate mass of a proton or neutron. Protons and neutrons make up almost all of the mass of an atom. Each electron has only about 1/1800 of the mass of a proton or neutron.

The **atomic mass** of an atom is a number that indicates how massive it is compared with another atom. This value is determined by adding the number of protons to the number of neutrons and expressing the result in atomic mass units or daltons.[1] The mass of the electrons is ignored because it is so small. The atomic mass number is indicated by a superscript to the left of the chemical symbol. The common form of the oxygen atom, with eight protons and eight neutrons in its nu-

[1]Unlike weight, mass is independent of the force of gravity. For convenience, however, we will consider mass and weight to be equivalent. Atomic weight has the same numerical value as atomic mass, but has no units.

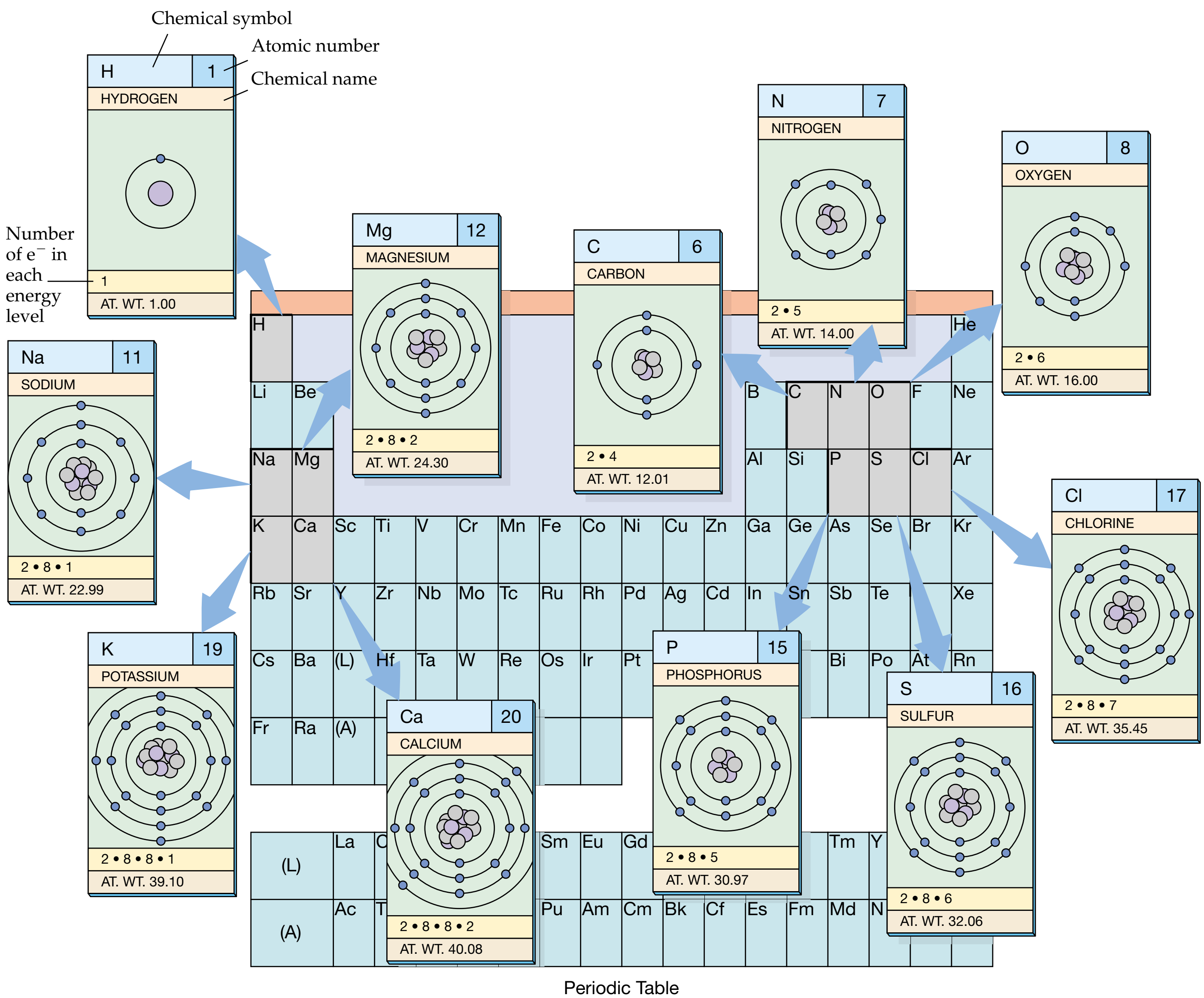


Figure 2–1 **The periodic table.** Note the Bohr models depicting the electron configuration of some biologically important atoms. Although the Bohr model does not depict electron configurations accurately, it is commonly used because of its simplicity and convenience.

cleus, has an atomic number of 8 and a mass of 16 atomic mass units. It is indicated by the symbol $^{16}_{8}O$.

The characteristics of protons, electrons, and neutrons are summarized below:

Particle	Charge	Approximate Mass	Location
Proton	positive	1 amu	nucleus
Neutron	neutral	1 amu	nucleus
Electron	negative	approx. 1/1800 amu	outside nucleus

Isotopes differ in number of neutrons

Most elements consist of a mixture of atoms with different numbers of neutrons and thus different masses. Such atoms are called **isotopes.** Isotopes of the same element have the same number of protons and electrons; only the number of neutrons varies. The three isotopes of hydrogen, $^{1}_{1}H$ (ordinary hydrogen), $^{2}_{1}H$ (deuterium), and $^{3}_{1}H$ (tritium), contain zero, one, and two neutrons, respectively. Bohr models of two isotopes of carbon, $^{12}_{6}C$ and $^{14}_{6}C$, are illustrated in Figure 2–2. The mass of an element is expressed as an average of the masses of its isotopes (weighted by their relative abundance in nature). For example,

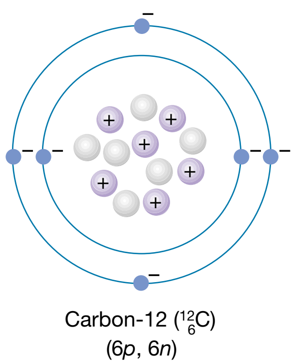


Figure 2–2 Isotopes differ in atomic mass. Carbon-12 ($^{12}_{6}C$) is the most common isotope of carbon. Its nucleus contains six protons and six neutrons, so its atomic mass is 12. Carbon-14 ($^{14}_{6}C$) is a rare radioactive carbon isotope. Because it contains eight neutrons, its atomic mass is 14.

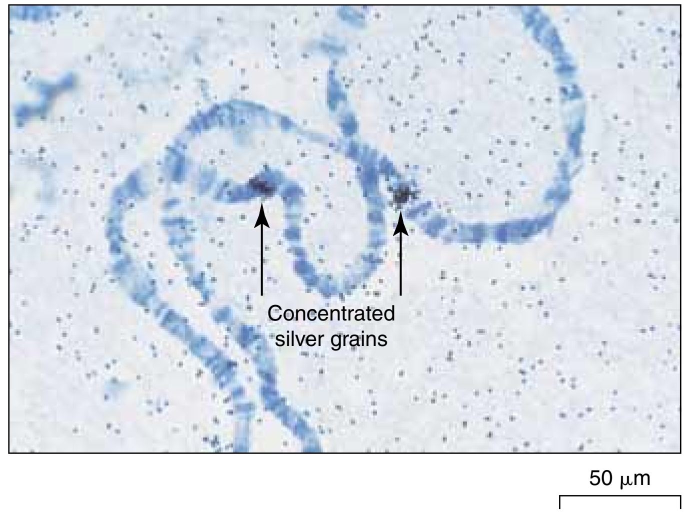


Figure 2–3 Autoradiography. The chromosomes of the fruit fly, *Drosophila melanogaster*, shown in this light micrograph have been covered with a special type of x-ray film in which silver grains (*dark spots*) are produced when tritium (^{3}H) undergoes radioactive decay. The concentrations of silver grains (*arrows*) mark the locations of specific DNA molecules. *©(Peter J. Bryant/Biological Photo Service)*

the atomic mass of hydrogen is not 1.0 amu, but 1.0079 amu, reflecting the natural occurrence of small amounts of deuterium and tritium.

All isotopes of a given element have essentially the same chemical characteristics. However, some isotopes are unstable and tend to break down, or decay, to a more stable isotope (usually becoming a different element). For example, the radioactive decay of $^{14}_{6}C$ yields the common form of nitrogen, $^{14}_{7}N$. Such isotopes are termed **radioisotopes** because they emit radiation when they decay. (The radioactive decay of $^{14}_{6}C$ occurs as a neutron decomposes to form a proton and a fast-moving electron, which is emitted from the atom as a form of radiation known as a β particle.) Sophisticated instruments allow scientists to detect and measure this and other types of radiation. Radioactive decay can also be detected by a method known as **autoradiography,** in which radiation causes the appearance of dark silver grains in special x-ray film (Fig. 2–3).

Despite the difference in the number of neutrons, the different isotopes of a given element are usually metabolized and/or localized in the organism in a similar way. For this reason, radioisotopes such as ^{3}H (tritium), ^{14}C, and ^{32}P have been extremely valuable research tools in studies ranging from determining the age of fossils, to DNA synthesis, to sugar transport in plants.

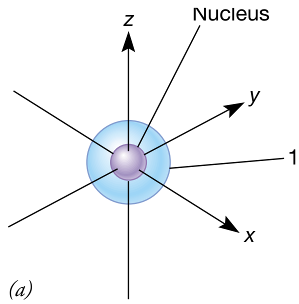
In medicine, radioisotopes are used for both diagnosis and treatment. The location and/or metabolism of a sugar, hormone, or drug can be followed in the body by labeling the substance with a radioisotope such as carbon-14 or tritium. For example, the active component in marijuana (tetrahydrocannabinol, or THC) can be labeled and administered intravenously. Then the amount of radioactivity in the blood and urine can be measured at successive intervals. Results of such measurements have determined that for several weeks this compound remains in the blood, and products of its metabolism can be detected in the urine. Radioisotopes are also used to test thyroid gland function, to measure the rate of red blood cell production, and to study many other aspects of body function and chemistry. Because radiation can interfere with cell division, radioisotopes have been used in the treatment of cancer (a disease often characterized by rapidly dividing cells).

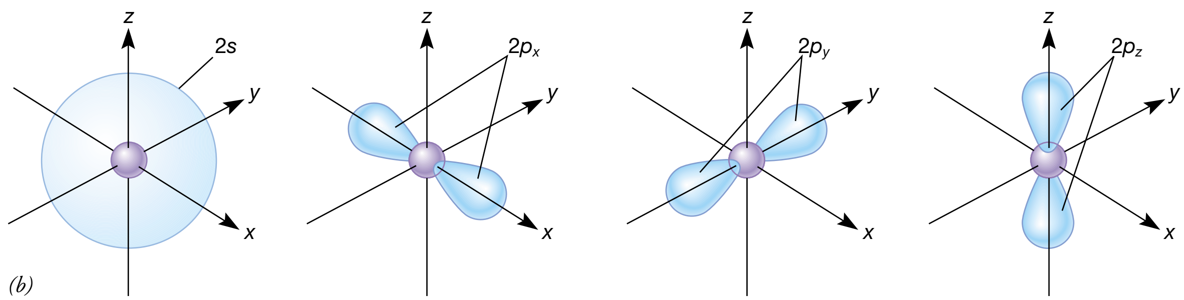
Electrons occupy orbitals corresponding to energy levels

Electrons move through characteristic regions of three-dimensional space, termed **orbitals.** Each orbital contains a maximum of two electrons. Because it is impossible to know an electron's position at any given time, orbitals are most accurately depicted as "electron clouds," shaded areas whose density is proportional to the probability that an electron is present there at any given instant. The energy of an electron depends on the orbital it occupies. Electrons in orbitals with similar energies, said to be at the same **principal energy level,** make up an **electron shell.** These are illustrated in Figure 2–4.

In general, electrons in a shell distant from the nucleus have greater energy than those in a shell close to the nucleus. This is because energy is required to move a negatively charged electron farther away from the positively charged nucleus. The most energetic electrons, known as **valence electrons,** are said to occupy the **valence shell.** The valence shell is represented as the outermost concentric ring in a Bohr model.

An electron can move to an orbital farther from the nucleus by receiving more energy, or it can give up energy and sink to a lower energy level in an orbital nearer the nucleus. Changes in electron energy levels are important in energy conversions in organisms. For example, in photosynthesis (see Chapter 8) light energy absorbed by chlorophyll molecules causes electrons to move to a higher energy level.





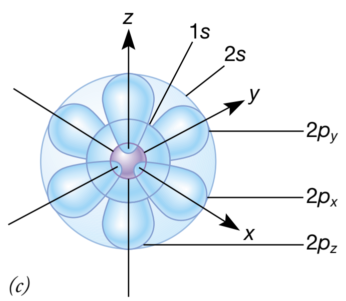


Figure 2–4 **Atomic orbitals.** Each orbital is represented as an "electron cloud." ***(a)*** The first principal energy level contains a maximum of two electrons, occupying a single spherical orbital (designated 1*s*). The electrons depicted in the diagram could be present anywhere within the deep blue area. ***(b)*** The second principal energy level includes four orbitals, each with a maximum of two electrons: one spherical (2*s*), and three dumbbell-shaped (2*p*) orbitals at right angles to each other. ***(c)*** Orbitals of the first and second principal energy levels are shown superimposed.

Atoms undergo chemical reactions

The chemical behavior of an atom is determined primarily by the number and arrangement of the valence electrons. The valence shell of hydrogen or helium is full (i.e., stable) when it contains two electrons. The valence shell of any other atom is full when it contains eight electrons. When the valence shell is not full, the atom tends to lose, gain, or share electrons to achieve a full outer shell. (The valence shells of all isotopes of an element are the same; this is why they have similar chemical properties and can substitute for each other in chemical reactions.)

Elements that fall into the same vertical column (said to belong to the same *group*) of the periodic table have similar chemical properties because their valence shells have similar tendencies to lose, gain, or share electrons. For example, chlorine and bromine, included in a group commonly known as the halogens, are highly reactive. Because their valence shells have seven electrons, they tend to gain an electron in chemical reactions. By contrast, hydrogen, sodium, and potassium each have a single valence electron, which they tend to give up or share with another atom. Helium (He) and neon (Ne) belong to a group referred to as the "noble gases." They are quite unreactive because their valence shells are full. Note the valence shells of some the elements important in organisms, including carbon, hydrogen, oxygen, and nitrogen, in Figure 2–1.

Atoms form molecules and compounds

Two or more atoms may combine chemically to form units called **molecules.** For example, when two atoms of oxygen combine chemically, a molecule of oxygen is formed. Atoms of *different* elements can combine to form chemical compounds. A **chemical compound** consists of two or more different elements combined in a fixed ratio. For example, water is a chemical compound consisting of hydrogen and oxygen in a ratio of 2:1. Water happens to be a molecular compound, with each molecule consisting of two atoms of hydrogen and one of oxygen. However, as we shall see, not all compounds are made up of molecules.

The properties of a chemical compound can be quite different from those of its component elements: at room temperature, water is usually a liquid, whereas hydrogen and oxygen are gases.

A substance can be described by a chemical formula

A **chemical formula** is a shorthand expression that describes the chemical composition of a substance. Chemical symbols indicate the types of atoms present, and subscript numbers indicate the ratios among the atoms. There are several types of chemical formulas, each providing specific kinds of information.

In a **molecular formula,** the subscripts indicate the actual numbers of each type of atom in a molecule. The formula for molecular oxygen, O_2, tells us that this molecule consists of two atoms of oxygen. The molecular formula for water, H_2O, indicates that each molecule consists of two atoms of hydrogen and one atom of oxygen. (Note that when a single atom of one type is present, the subscript number 1 is never written.)

Another type of formula is the **structural formula,** which shows not only the types and numbers of atoms in a molecule, but also their arrangement. From the molecular formula for water, H_2O, you would not know whether the atoms were arranged H—H—O or H—O—H. The structural formula, H—O—H, settles the matter, indicating that the two hydrogen atoms are attached to the oxygen atom.

One mole of any substance contains the same number of units

The **molecular mass** of a compound is the sum of the atomic masses of the component atoms of a single molecule; thus, the molecular mass of water, H_2O, is $(2 \times 1 \text{ amu}) + (16 \text{ amu})$, or 18 amu. (Owing to the presence of isotopes, atomic mass values are not whole numbers. However, for easy calculation each atomic mass value has been rounded off to a whole number.) The molecular mass of glucose ($C_6H_{12}O_6$), a simple sugar that is a key compound in cellular metabolism, is $(6 \times 12 \text{ amu}) + (12 \times 1 \text{ amu}) + (6 \times 16 \text{ amu})$, or 180 amu.

The amount of an element or compound whose mass in grams is equivalent to its atomic or molecular mass is termed 1 **mole.** Thus 1 mole of glucose has a mass of 180 grams. The mole is a useful concept because it allows us to make meaningful comparisons between atoms and molecules of very different mass. This is because *one mole of any substance always has exactly the same number of units,* whether they be small atoms or large molecules. The very large number of units in a mole, 6.02×10^{23}, is known as **Avogadro's number,** named for the Italian physicist, Amadeo Avogadro, who first calculated it. Thus 1 mole (180 grams) of glucose contains 6.02×10^{23} molecules, as does 1 mole (2 grams) of molecular hydrogen (H_2). Although it is impossible to count atoms and molecules individually, this fact allows a scientist to count them simply by weighing a sample. Molecular biologists usually deal with smaller values, either millimoles (a mmole is one thousandth of a mole) or micromoles (a μmole is one millionth of a mole).

The mole concept allows us to make useful comparisons among solutions. A 1-molar solution, represented by 1 *M,* contains 1 mole of that substance dissolved in 1 liter of solution. For example, we can compare 1 liter of a 1 *M* solution of glucose with 1 liter of a 1 *M* solution of sucrose (table sugar). They differ in the mass of the dissolved sugar (180 g and 340 g, respectively), but they each contain 6.02×10^{23} sugar molecules.

Chemical equations describe chemical reactions

During any moment in the life of an organism, be it a mushroom or a butterfly, many complex chemical reactions are taking place. Chemical reactions—for example, the reaction between glucose and oxygen—can be described by means of chemical equations:

$$\underset{\text{Glucose}}{C_6H_{12}O_6} + \underset{\text{Oxygen}}{6\,O_2} \longrightarrow \underset{\text{Carbon dioxide}}{6\,CO_2} + \underset{\text{Water}}{6\,H_2O} + \text{Energy}$$

In a chemical equation, the **reactants** (the substances that participate in the reaction) are generally written on the left side, and the **products** (the substances formed by the reaction) are written on the right side. The arrow means "yields" and indicates the direction in which the reaction tends to proceed.

Chemical compounds react with each other in quantitatively precise ways. The numbers preceding the chemical symbols or formulas (known as *coefficients*) indicate the relative number of atoms or molecules reacting. For example, 1 mole of glucose burned in a fire or metabolized in a cell reacts with 6 moles of oxygen to form 6 moles of carbon dioxide and 6 moles of water. Many reactions can proceed in the reverse direction (to the left) as well as in the forward direction (to the right); at **equilibrium** the rates of the forward and reverse reactions are equal (see Chapter 6). Reversible reactions are indicated by double arrows:

$$\underset{\text{Carbon dioxide}}{CO_2} + \underset{\text{Water}}{H_2O} \rightleftharpoons \underset{\text{Carbonic acid}}{H_2CO_3}$$

In this example, the arrows are drawn different lengths to indicate that when the reaction reaches equilibrium there will be more reactants (CO_2 and H_2O) than product (H_2CO_3).

Atoms are joined by chemical bonds

The atoms of a compound are held together by forces of attraction called **chemical bonds.** Each bond represents a certain amount of chemical energy. **Bond energy** is the energy necessary to break a bond. The valence electrons dictate how many bonds an atom can participate in. The two principal types of strong chemical bonds are covalent bonds and ionic bonds.

In covalent bonds electrons are shared

Covalent bonds involve the sharing of electrons between atoms in a way that results in each having a filled valence shell. A compound consisting mainly of covalent bonds is called a **covalent compound.** A simple example of a covalent bond is the joining of two hydrogen atoms in a molecule of hydrogen gas, H_2. Each atom of hydrogen has one electron, but two electrons are required to complete its valence shell. The hydrogen atoms have equal capacities to attract electrons, so neither donates an electron to the other. Instead, the two hydrogen atoms share their single electrons so that each of the two electrons is attracted simultaneously to the two protons in the two hydrogen nuclei. The two electrons thus whirl around *both* atomic nuclei, joining the two atoms.

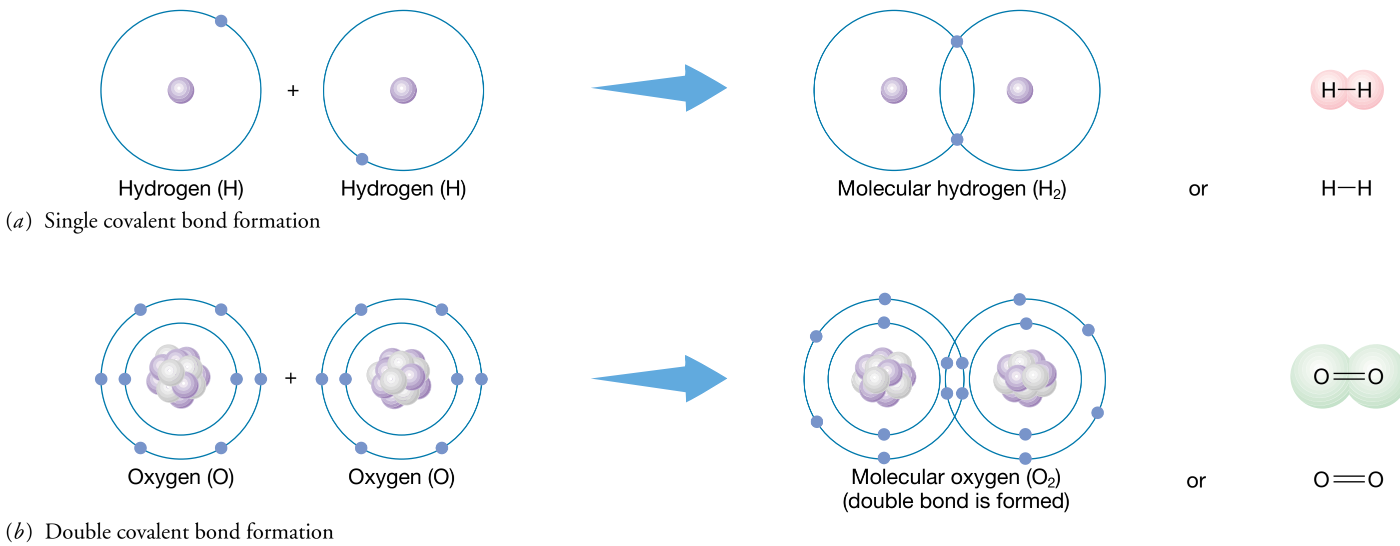


Figure 2–5 **Electron sharing in covalent compounds.** ***(a)*** Two hydrogen atoms achieve stability by sharing electrons, thereby forming a molecule of hydrogen. In the structural formula shown on the right, the straight line between the hydrogen atoms represents a single covalent bond. ***(b)*** In molecular oxygen, two oxygen atoms share two pairs of electrons, forming a double covalent bond.

A simple way of representing the electrons in the valence shell of an atom is to use dots placed around the chemical symbol of the element. Such a representation is called the *Lewis structure* of the atom. In a water molecule, two hydrogen atoms are covalently bonded to an oxygen atom:

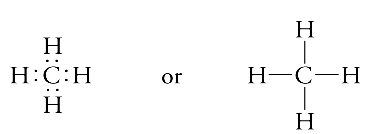
$$\mathrm{H\cdot + H\cdot + \cdot\ddot{\underset{\cdot\cdot}{O}}\cdot \longrightarrow H:\ddot{\underset{\cdot\cdot}{O}}:H}$$

Oxygen has six valence electrons; by sharing electrons with two hydrogen atoms, it completes its valence shell of eight. At the same time each hydrogen atom obtains a complete valence shell of two. (Note that in the structural formula H—O—H, each pair of shared electrons constitutes a covalent bond, represented by a solid line. Unshared electrons are usually omitted in a structural formula.)

The carbon atom has four electrons in its valence shell. These four electrons are available for covalent bonding:

$$\mathrm{\cdot\dot{\underset{\cdot}{C}}\cdot}$$

When one carbon and four hydrogen atoms share electrons, a molecule of methane, CH_4, is formed:

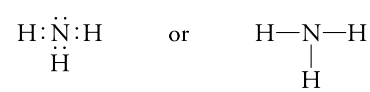


Lewis structure Structural formula

The nitrogen atom has five electrons in its valence shell. Recall that each orbital can hold a maximum of two electrons. Usually two electrons occupy one orbital, leaving three available for sharing with other atoms:

$$\mathrm{\cdot\ddot{\underset{\cdot}{N}}\cdot}$$

When a nitrogen atom shares electrons with three hydrogen atoms, a molecule of ammonia, NH_3, is formed:



Lewis structure Structural formula

When one pair of electrons is shared between two atoms, the covalent bond is referred to as a **single covalent bond** (Fig. 2–5*a*). Two oxygen atoms may achieve stability by forming covalent bonds with one another. Each oxygen atom has six electrons in its outer shell. To become stable, the two atoms share two pairs of electrons, forming molecular oxygen (Fig. 2–5*b*). When two pairs of electrons are shared in this way, the covalent bond is called a **double covalent bond,** which is represented by two parallel solid lines. Similarly, a **triple covalent bond** (represented by three parallel solid lines) is formed when three pairs of electrons are shared between two atoms.

The number of covalent bonds usually formed by the atoms commonly present in biologically important molecules is summarized as follows:

Atom	Symbol	Covalent Bonds
Hydrogen	H	1
Oxygen	O	2
Carbon	C	4
Nitrogen	N	3
Phosphorus	P	5
Sulfur	S	2

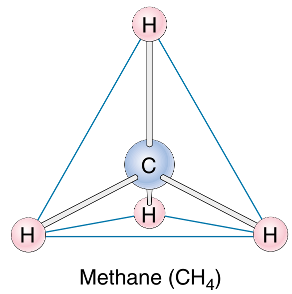


Figure 2–6 **Methane.** The four hydrogens are located at the corners of a tetrahedron.

The function of a molecule is related to its shape

Each kind of molecule has a characteristic size and a general overall shape. Although the shape of a molecule may change (within certain limits), the functions of molecules in living cells are largely dictated by their geometric shapes. A molecule that consists of two atoms, for example, is linear. Molecules composed of more than two atoms may have more complicated shapes. The geometric shape of a molecule provides the optimal distance between the atoms to counteract the repulsion of electron pairs.

When an atom forms covalent bonds with other atoms, the orbitals in the valence shell may become rearranged, thereby affecting the shape of the resulting molecule. For example, when four hydrogen atoms combine with a carbon atom to form a molecule of methane (CH_4), the valence shell orbitals of the carbon become rearranged such that a geometric structure known as a *tetrahedron* is formed, with one hydrogen atom present at each of its four corners (Fig. 2–6).

Covalent bonds can be nonpolar or polar

The atoms of each element have a characteristic affinity for electrons. **Electronegativity** is a measure of an atom's attraction for electrons in chemical bonds. Very electronegative atoms are sometimes called "electron-greedy." When covalently bound atoms have similar electronegativities, the electrons are shared equally, and the covalent bond is described as **nonpolar.** The covalent bond of the hydrogen molecule is nonpolar, as are the covalent bonds of molecular oxygen and methane.

In a covalent bond between two different elements, such as oxygen and hydrogen, the electronegativities of the atoms may be different. If so, electrons are pulled closer to the atomic nucleus of the element with the greater electron affinity (in this case, oxygen). A covalent bond between atoms that differ in electronegativity is called a **polar covalent bond.** Such a bond has two dissimilar ends (or poles), one with a partial positive charge and the other with a partial negative charge. Each of the two covalent bonds in water is polar because there is a partial positive charge at the hydrogen end of the bond and a partial negative charge at the oxygen end, where the "shared" electrons are more likely to be found (Fig. 2–7).

Covalent bonds differ in their degree of polarity, ranging from those in which the electrons are exactly shared (as in the nonpolar hydrogen molecule) to those in which the electrons are much closer to one atom than to the other (as in water). Oxygen is quite electronegative and forms polar covalent bonds with carbon, hydrogen, and many other atoms. Nitrogen is also relatively electronegative, although less so than oxygen.

A molecule with one or more polar covalent bonds can be polar even though it is electrically neutral as a whole. A **polar molecule** has one end with a partial positive charge and another end with a partial negative charge. One example is water (Fig. 2–7). The polar bonds between the hydrogens and the oxygens are arranged in a "V" shape, rather than linearly. The oxygen end therefore constitutes the negative pole, and the end with the two hydrogens is the positive pole.

Ionic bonds form between cations and anions

Some atoms or groups of atoms are not electrically neutral. A particle with one or more units of electrical charge is called an **ion.** An atom becomes an ion if it gains or loses one or more electrons. An atom with one, two, or three electrons in its va-

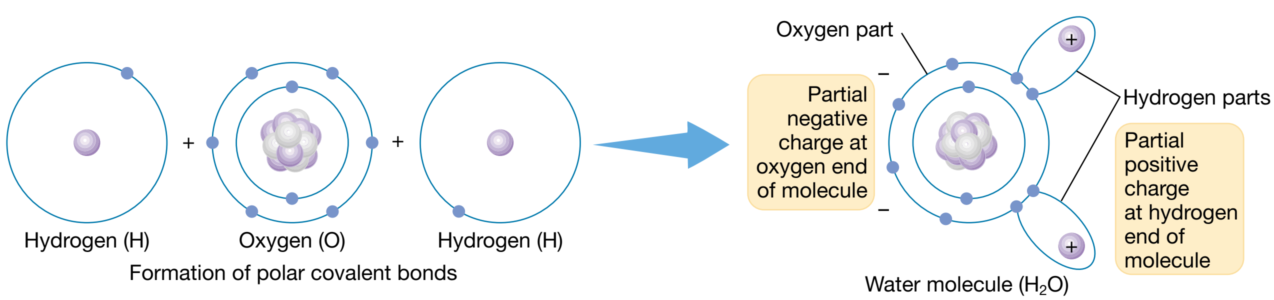


Figure 2–7 **Water, a polar molecule.** Note that the electrons tend to stay closer to the nucleus of the oxygen atom than to the hydrogen nuclei. This results in a partial negative charge on the oxygen portion of the molecule and a partial positive charge at the hydrogen end. Although the water molecule as a whole is electrically neutral, it is a polar covalent compound.

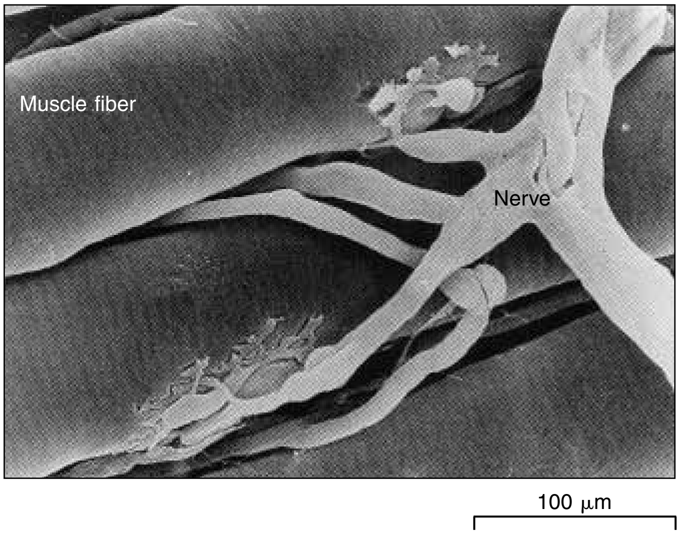


Figure 2–8 **Ions and biological processes.** Sodium, potassium, and chloride ions are essential for this nerve cell to stimulate these muscle fibers. Calcium ions in the muscle cell are required for muscle contraction. *(D.W. Fawcett)*

lence shell tends to lose electrons to other atoms. Such an atom then becomes positively charged because its nucleus contains more protons than the number of electrons orbiting around the nucleus. These positively charged ions are termed **cations.** Atoms with five, six, or seven valence electrons tend to gain electrons from other atoms and become negatively charged **anions.**

The properties of ions are very different from those of the electrically neutral atoms from which they were derived. For example, although chlorine gas is a poison, chloride ions (Cl^-) are essential to life. Because their electrical charges provide a basis for many interactions, cations and anions are involved in energy transformations within the cell, the transmission of nerve impulses, muscle contraction, and many other life processes (Fig. 2–8).

A group of atoms can also become an ion (polyatomic ion). Unlike a single atom, a group of atoms can lose or gain protons (derived from hydrogen atoms) as well as electrons. Therefore, a group of atoms can become a cation if it loses one or more electrons or gains one or more protons. A group of atoms becomes an anion if it gains one or more electrons or loses one or more protons.

An **ionic bond** forms as a consequence of the attraction between the positive charge of a cation and the negative charge of an anion. An **ionic compound** is a substance consisting of anions and cations bonded together by their opposite charges.

A good example of how ionic bonds are formed is the attraction between sodium ions and chloride ions. A sodium atom has one electron in its valence shell. It cannot fill its valence shell by obtaining seven electrons from other atoms, for it would then have a very large unbalanced negative charge. Instead, it gives up its single valence electron to a very electronegative atom, such as chlorine, which acts as an electron acceptor (Fig. 2–9). Chlorine cannot give up the seven electrons in its valence shell, because it would then have a vast positive charge. Instead it strips an electron from an electron donor (sodium in this example) to complete its valence shell.

When sodium reacts with chlorine, its valence electron is transferred completely to chlorine. Sodium is now a cation, with one unit of positive charge (Na^+). Chlorine is now an anion, a chloride ion with one unit of negative charge (Cl^-).

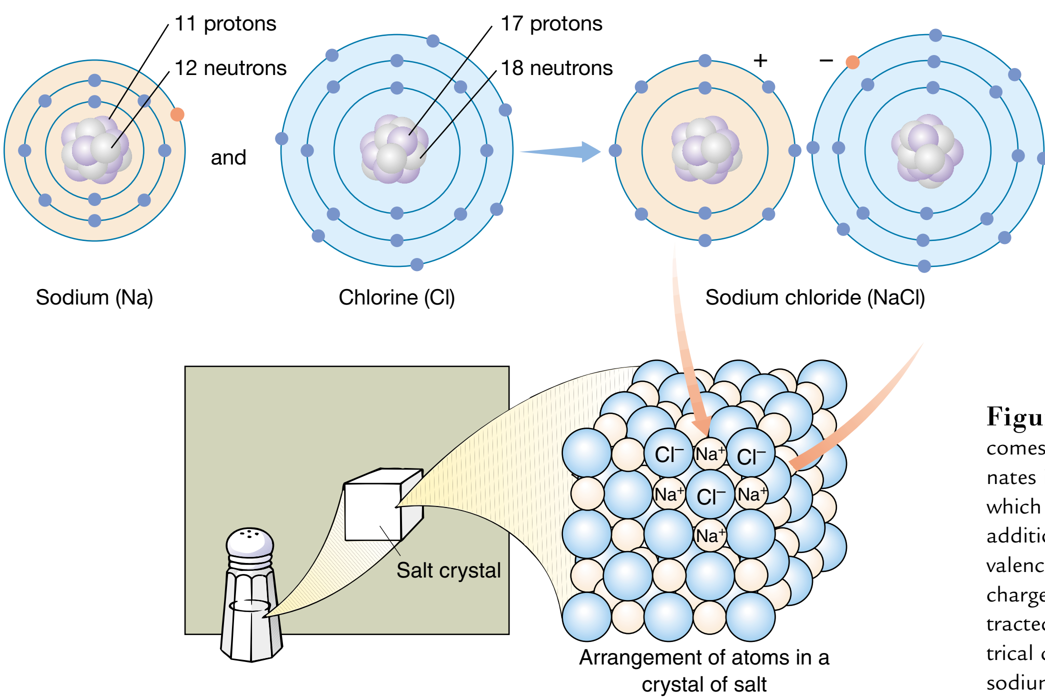


Figure 2–9 **Ionic bonding.** Sodium becomes a positively charged ion when it donates its single valence electron to chlorine, which has seven valence electrons. With this additional electron, chlorine completes its valence shell and becomes a negatively charged chloride ion. These ions are attracted to one another by their unlike electrical charges, forming the ionic compound sodium chloride.

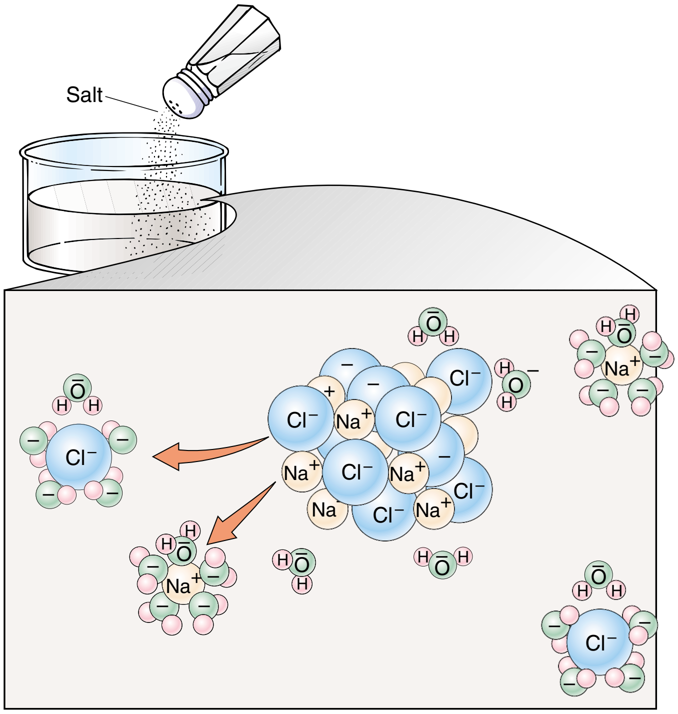


Figure 2–10 Hydration of an ionic compound. When the crystal of NaCl is added to water, the sodium and chloride ions are pulled apart as the partial negative ends of the water molecules are attracted to the positive sodium ions, and the partial positive ends of the water molecules are attracted to the negative chloride ions. When the NaCl is dissolved, each Na^+ and Cl^- is surrounded by water molecules electrically attracted to it.

These ions attract each other as a result of their opposite charges. They are held together by this electrical attraction in ionic bonds to form NaCl, sodium chloride,[2] or common table salt.

The term *molecule* does not adequately explain the properties of ionic compounds such as NaCl. When NaCl is in its solid crystal state, each ion is actually surrounded by six ions of opposite charge. The molecular formula NaCl indicates that sodium ions and chloride ions are present in a one-to-one ratio, but in the actual crystal, no discrete molecules composed of one Na^+ and one Cl^- ion are present.

Compounds joined by ionic bonds, such as sodium chloride, have a tendency to *dissociate* (separate) into their individual ions when placed in water:

$$\underset{\text{Sodium chloride}}{NaCl} \xrightarrow{\text{in } H_2O} \underset{\text{Sodium ion}}{Na^+} + \underset{\text{Chloride ion}}{Cl^-}$$

[2]In both covalent and ionic binary compounds (*binary* denotes compounds consisting of two elements), the element having the greater attraction for electrons is named second, and an-*ide* ending is added to the stem name, e.g., sodium chloride, hydrogen fluoride. The -*ide* ending is also used to indicate an anion, as in chloride (Cl^-) and hydroxide (OH^-).

In the solid form of an ionic compound, the ionic bonds are very strong. Water, however, is an excellent **solvent;** as a liquid it is capable of dissolving many substances, particularly those that are polar or ionic. This is because of the polarity of water molecules. The localized partial positive charges (on the hydrogen atoms) and partial negative charges (on the oxygen atom) on each water molecule attract the anions and cations on the surface of an ionic solid. As a result, the solid dissolves. A dissolved substance is referred to as a **solute.** In solution, each cation and anion of the ionic compound is surrounded by oppositely charged ends of the water molecules (Fig. 2–10). This process is known as **hydration.** Hydrated ions still interact with each other to some extent, but the transient ionic bonds formed are much weaker than those in a solid crystal.

Hydrogen bonds are weak attractions

Another type of bond important in organisms is the **hydrogen bond.** When hydrogen combines with oxygen (or with another relatively electronegative atom such as nitrogen), it has a partial positive charge because its electron spends more time closer to the electronegative atom. Hydrogen bonds tend to form between an atom with a partial negative charge and a hydrogen atom that is covalently bonded to oxygen or nitrogen (Fig. 2–11). The atoms involved may be in two parts of the same large molecule or in two different molecules. Water molecules interact with each other extensively through hydrogen bond formation.

Hydrogen bonds are readily formed and broken. Although individually relatively weak, hydrogen bonds are collectively strong when present in large numbers. Furthermore, they have a specific length and orientation. As we will see in Chapter 3, these features are very important in helping determine the three-dimensional structure of large molecules such as DNA and proteins.

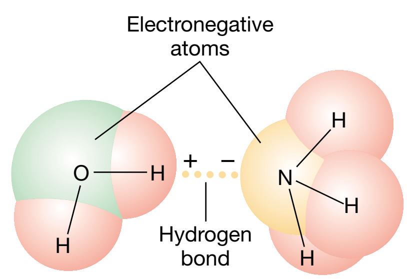


Figure 2–11 Hydrogen bonding. A hydrogen atom in a water molecule has a partial positive charge because of its polar covalent bond with oxygen. Nitrogen is relatively electronegative and, in molecules like ammonia (NH_3), has a partial negative charge because of its polar covalent bonds with hydrogen. Here, the nitrogen atom of the ammonia molecule is joined by a hydrogen bond to a hydrogen atom of a water molecule. A hydrogen bond is generally indicated by a dotted line.

Electrons and their energy are transferred in redox reactions

Many of the energy conversions that go on in a cell involve reactions in which an electron is transferred from one substance to another. This is because the transfer of an electron also involves the transfer of the energy of that electron. Such an electron transfer is called an oxidation-reduction, or **redox reaction.** Both cellular respiration (Chapter 7) and photosynthesis (Chapter 8) are essentially redox processes.

Rusting, which is the combination of iron (symbol Fe) with oxygen, is a simple illustration of oxidation and reduction:

$$4\ Fe + 3\ O_2 \longrightarrow \underset{\text{Iron (III) oxide}}{2\ Fe_2O_3}$$

Oxidation and reduction always occur together, but initially we will discuss them separately. **Oxidation** is a chemical process in which an atom, ion, or molecule loses electrons. In rusting, each iron atom becomes oxidized as it loses three electrons.

$$4\ Fe \longrightarrow 4\ Fe^{3+} + 12\ e^-$$

The e^- is a symbol for an electron; the + superscript represents an electron deficit. (When an atom loses an electron, it acquires one unit of positive charge from the excess of one proton. In our example, each iron atom loses three electrons and acquires three units of positive charge.)

You will recall that the oxygen atom is very electronegative, able to remove electrons from other atoms. In this reaction, oxygen gains electrons from iron.

$$3\ O_2 + 12\ e^- \longrightarrow 6\ O^{2-}$$

Oxygen becomes reduced when it accepts electrons from the iron. **Reduction** is a chemical process in which an atom, ion, or molecule *gains* electrons. (The term reduction refers to the fact that the gain of an electron results in the *reduction* of any positive charge that might be present.)

Redox reactions occur simultaneously because one substance must accept the electrons that are removed from the other. In a redox reaction, one component, the *oxidizing agent,* accepts one or more electrons and becomes reduced. Oxidizing agents other than oxygen are known, but oxygen is such a common one that its name was given to the process. Another reaction component, the *reducing agent,* gives up one or more electrons and becomes oxidized.

In our example there was a complete transfer of electrons from iron to oxygen. Some redox reactions are not so obvious, however. In these, electrons simply move farther from the reducing agent and closer to the oxidizing agent.

Electrons are not easily removed from covalent compounds unless an entire atom is removed. In cells, oxidation often involves the removal of a hydrogen *atom* (an electron plus a proton that "goes along for the ride") from a compound; reduction often involves the addition of hydrogen (see Chapter 6).

Water is essential to life

A large part of the mass of most organisms is water. In human tissues the percentage of water ranges from 20% in bones to 85% in brain cells. About 70% of our total body weight is water; as much as 95% of a jellyfish and certain plants is water. Water is the source, through photosynthesis (see Chapter 8), of the oxygen in the air we breathe, and its hydrogen atoms become incorporated into many organic compounds. Water is also the solvent for most biological reactions and a reactant or product in many chemical reactions.

Water is not only important inside organisms but also is one of the principal environmental factors affecting them. Many organisms live within the ocean or in freshwater rivers, lakes, or puddles. Water's unique combination of physical and chemical properties has permitted living things to originate, survive, and evolve on Earth (Fig. 2–12).

Water molecules are polar

As discussed previously, water molecules are polar, that is, one end of each molecule bears a partial positive charge and the other a partial negative charge (see Fig. 2–7). The water molecules in liquid water and in ice associate by hydrogen bonds. The hydrogen atom of one water molecule, with its partial positive charge, is attracted to the oxygen atom of a neighboring water molecule, with its partial negative charge, forming a hydrogen bond. An oxygen in a water molecule has two regions of partial negative charge, and each of the two hydrogens has a partial positive charge. Each water molecule can therefore form hydrogen bonds with a maximum of four neighboring water molecules (Fig. 2–13).

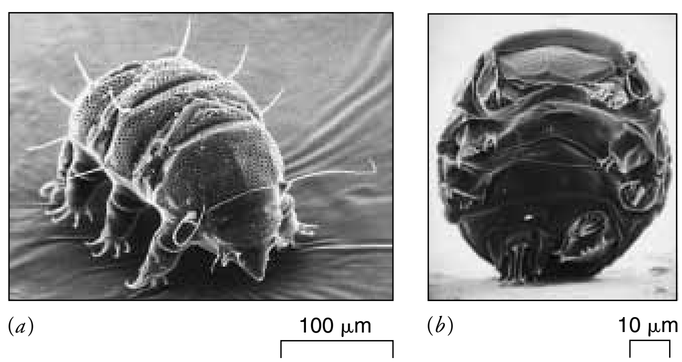


Figure 2–12 **Tardigrade.** ***(a)*** Commonly known as "water bears," tardigrades such as these members of the genus *Echiniscus* are small (less than 1.2 mm long) animals that normally live in moist habitats, such as thin films of water on mosses. ***(b)*** When subjected to dessication tardigrades assume a barrel-shaped form known as a *tun*, remaining in this state, motionless but alive, for as long as 100 years. When rehydrated they assume their normal appearance and activities. *(a, Diane R. Nelson; b, Robert O. Schuster, courtesy of Diane R. Nelson)*

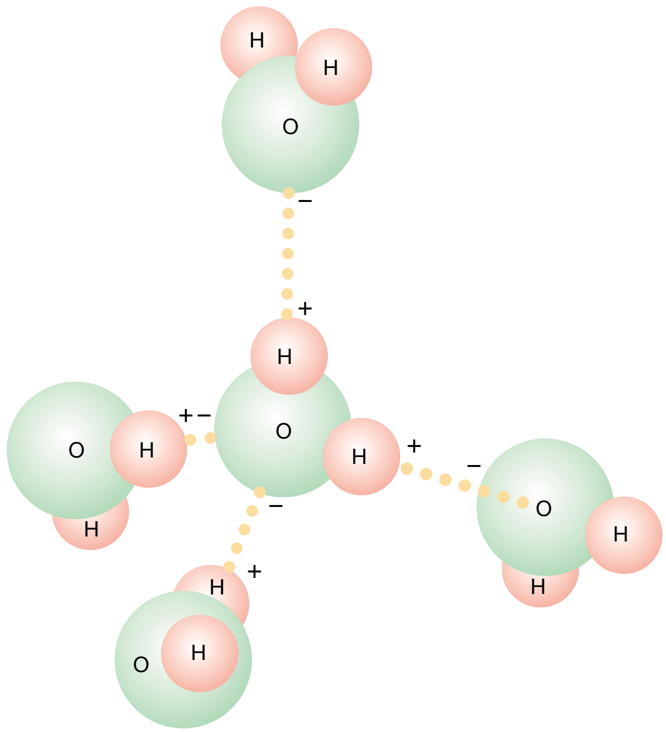


Figure 2–13 **Hydrogen bonding of water molecules.** Each water molecule can form hydrogen bonds (dotted lines) with as many as four neighboring water molecules.

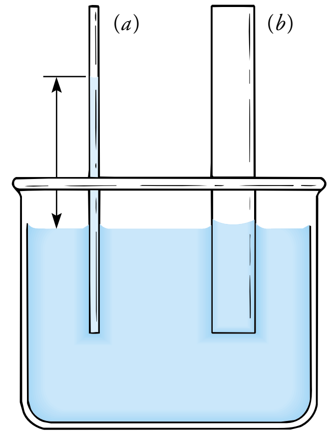


Figure 2–14 **Capillary action.** ***(a)*** In a narrow tube, adhesive forces attract water molecules to the glass wall of the tube. Other water molecules inside the tube are then "pulled along" by cohesive forces, which are actually due to hydrogen bonds between the water molecules. ***(b)*** In the wider tube, a smaller percentage of the water molecules line the glass wall. Because of this, the adhesive forces are not strong enough to overcome the cohesive forces of the water beneath the surface level of the container, and water in the tube rises only slightly.

Water is the principal solvent in organisms

Because its molecules are polar, water is an excellent solvent, a liquid capable of dissolving many different kinds of substances, especially polar and ionic compounds. Previously in this chapter, we discussed how polar water molecules pull the ions of ionic compounds apart so that they dissociate (see Fig. 2–10). Because of its solvent properties and the tendency of the atoms in certain compounds to form ions when in solution, water plays an important role in facilitating chemical reactions. Substances that interact readily with water are said to be **hydrophilic** ("water-loving"). Examples include table sugar (a polar compound) and NaCl (an ionic compound), which dissolve readily in water. Not all substances in organisms are hydrophilic, however. Many **hydrophobic** ("water-fearing") substances found in living things are especially important because of their ability to form structures that are not dissolved by water. Examples, to be discussed more fully in Chapter 3, include fats and other nonpolar substances.

Hydrogen bonding makes water cohesive and adhesive

Water molecules have a very strong tendency to stick to each other; that is, they are **cohesive.** This is due to the hydrogen bonds among the molecules. Water molecules also stick to many other kinds of substances, most notably those with charged groups of atoms or molecules on their surfaces. These **adhesive** forces explain how water makes things wet.

A combination of adhesive and cohesive forces accounts for the tendency, termed **capillary action,** of water to move in narrow tubes, even against the force of gravity (Fig. 2–14). For example, water moves through the microscopic spaces between soil particles to the roots of plants by capillary action. Because of the cohesive nature of water molecules, any force exerted on part of a column of water will be transmitted to the column as a whole. The major mechanism of water movement in plants (see Chapter 33) depends on this fact.

Water has a high degree of **surface tension** because of the cohesiveness of its molecules, which have a much greater attraction for each other than for molecules in the air. Thus, water molecules at the surface crowd together, producing a strong layer as they are pulled downward by the attraction of other water molecules beneath them (Fig. 2–15).

Water helps maintain a stable temperature

Raising the temperature of a substance involves adding heat energy to make its molecules move faster, that is, to increase the **kinetic energy** (energy of motion) of the molecules (see Chapter 6). The term **heat** refers to the *total* amount of kinetic energy in a sample of a substance; **temperature** refers to the *average* kinetic energy of the particles. Water has a high **specific heat;** that is, the amount of energy required to raise the temperature of water is quite large. A **calorie** is a unit of heat energy (equivalent to 4.184 joules) that equals the amount

Figure 2–15 **Surface tension of water.** Hydrogen bonding between water molecules is responsible for the surface tension of water, which is strong enough to support these water striders, and causes the dimpled appearance of the surface. Although they are denser than water, water striders can walk on the surface of a pond because fine hairs at the ends of their legs spread their weight over a large area. *(Dennis Drenner)*

of heat required to raise the temperature of 1 gram of water 1 degree Celsius. The specific heat of water is therefore 1 calorie per gram of water per degree Celsius. Most other common substances have much lower specific heat values.

The high specific heat of water results from the hydrogen bonding of its molecules. Some of the hydrogen bonds holding the water molecules together must first be broken to permit the molecules to move more freely. Much of the energy added to the system is used up in breaking the hydrogen bonds, and only a portion of the heat energy is available to speed the movement of the water molecules (thereby increasing the temperature of the water). Conversely, when liquid water changes to ice, additional hydrogen bonds must be formed, liberating a great deal of heat into the environment.

Because so much heat input is required to raise the temperature of water (and so much heat is lost when the temperature is lowered), the ocean and other large bodies of water have relatively constant temperatures. Thus, many organisms living in the ocean are provided with a relatively constant environmental temperature. The properties of water are crucial in stabilizing temperatures on the surface of Earth. Although surface water is only a thin film relative to Earth's volume, the quantity is enormous compared to the exposed land mass. This relatively large mass of water resists both the warming effect of heat and the cooling effect of low temperatures. In addition, hydrogen bonding gives ice unique properties that have important environmental consequences (see *Making the Connection: Hydrogen Bonding and the Environment*).

The high water content of organisms helps them maintain relatively constant internal temperatures. Such minimizing of temperature fluctuations is important because biological reactions can take place only within a relatively narrow temperature range.

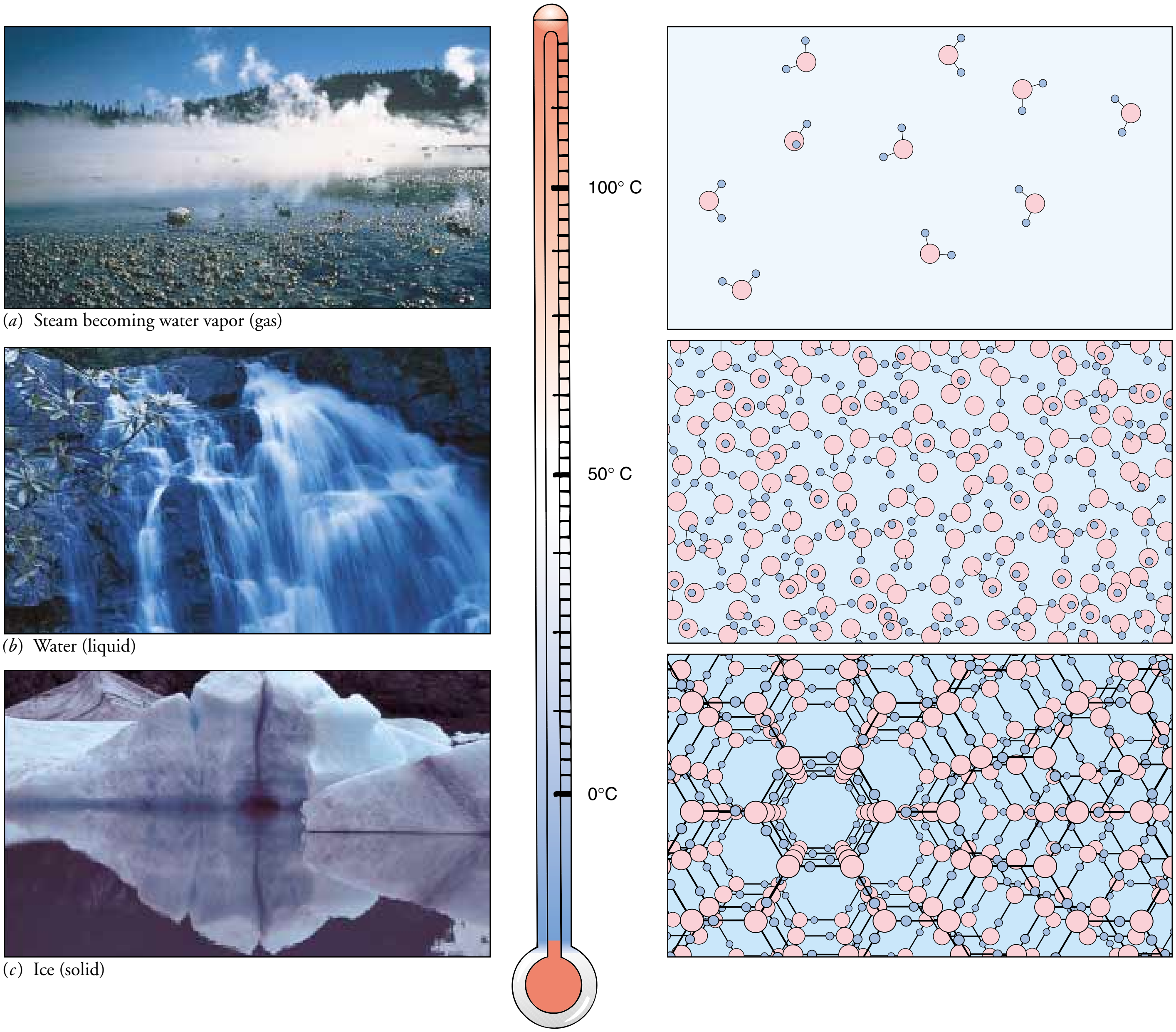
Because its molecules are held together by hydrogen bonds, water has a high **heat of vaporization.** To change 1 gram of liquid water into 1 gram of water vapor, 540 calories of heat are required. The heat of vaporization of most other common liquid substances is much less. As a sample of water is heated, some molecules are moving much faster than others (that is, they have more heat energy). These faster moving molecules are more likely to escape the liquid phase and enter the vapor phase (Fig. 2–16). When they do, they take their heat energy with them (thus lowering the temperature of the sample, in a process called **evaporative cooling**). For this reason the human body can dissipate excess heat as sweat evaporates from the skin, and a leaf can keep cool in the bright sunlight as water evaporates from its surface.

MAKING THE CONNECTION

Hydrogen Bonding and the Environment

Why does ice float? This is because hydrogen bonds contribute to another important property of water. Liquid water expands as it freezes because the hydrogen bonds joining the water molecules in the crystalline lattice keep the molecules far enough apart to give ice a density about 10% less than the density of liquid water (see Fig. 2–16). When ice has been heated enough to raise its temperature above 0° C (32° F), these hydrogen bonds among the water molecules are broken, freeing the molecules to slip closer together. The density of water is greatest at 4° C, above which water begins to expand again as the speed of its molecules increases. As a result, ice floats on the denser cold water.

This unusual property of water has been important in enabling life as we know it to appear, survive, and evolve on Earth. If ice had a greater density than water, it would sink; eventually all ponds, lakes, and even the ocean would freeze solid from the bottom to the surface, making life impossible. When a body of deep water cools, it becomes covered with floating ice. The ice insulates the liquid water below it, preventing it from freezing and permitting a variety of animals and plants to survive below the icy surface.



(*a*) Steam becoming water vapor (gas)

(*b*) Water (liquid)

(*c*) Ice (solid)

Figure 2–16 **Three forms of water.** ***(a)*** When water boils, as in this hot spring at Yellowstone National Park, many hydrogen bonds are broken, causing steam, consisting of minuscule water droplets, to form. If most of the remaining hydrogen bonds are subsequently broken, the molecules begin to move freely and rapidly as water vapor (a gas). ***(b)*** Water molecules in a liquid state continually form, break, and re-form hydrogen bonds with each other. ***(c)*** In ice, each water molecule participates in four hydrogen bonds with adjacent molecules, resulting in a regular, evenly distanced crystalline lattice structure. Because the water molecules move apart slightly as the hydrogen bonds form, water expands as it freezes; thus ice floats on water. *(a, Woodbridge Wilson/National Park Service; b, Gary R. Bonner; c, Barbara O'Donnell/Biological Photo Service)*

ACIDS ARE PROTON DONORS; BASES ARE PROTON ACCEPTORS

Water molecules have a slight tendency to **ionize,** that is, to dissociate into hydrogen ions (H^+) and hydroxide ions (OH^-).[3] In pure water, a very small number of water molecules ionize. This slight tendency of water to dissociate is reversible as hydrogen ions and hydroxide ions reunite to form water:

$$HOH \rightleftharpoons H^+ + OH^-$$

[3]The H^+ immediately combines with a negatively charged region of a water molecule, forming a hydronium ion (H_3O^+). However, by convention H^+, rather than the more accurate H_3O^+, is used.

Because each water molecule splits into one hydrogen ion and one hydroxide ion, the concentrations of hydrogen ions and hydroxide ions in pure water are exactly equal (0.0000001 or 10^{-7} moles per liter for each ion). Such a solution is said to be **neutral,** neither acidic nor basic (alkaline).

An **acid** is a substance that dissociates in solution to yield hydrogen ions (H^+) and an anion.

$$\text{Acid} \longrightarrow H^+ + \text{Anion}$$

An acid is a proton *donor.* (Recall that a hydrogen ion, or H^+, is nothing more than a proton.) An acidic solution has a hydrogen ion concentration that is higher than its hydroxide ion concentration. Acidic solutions turn blue litmus paper red and have a sour taste. Hydrochloric acid (HCl) and sulfuric acid (H_2SO_4) are examples of inorganic acids. Lactic acid ($CH_3CHOHCOOH$) from sour milk and acetic acid (CH_3COOH) from vinegar are two common organic acids.

A **base** is defined as a proton *acceptor.* Most bases are substances that dissociate to yield a hydroxide ion (OH^-) and a cation when dissolved in water. A hydroxide ion can act as a base by accepting a proton (H^+) to form water. Sodium hydroxide (NaOH) is a common inorganic base.

$$NaOH \longrightarrow Na^+ + OH^-$$
$$OH^- + H^+ \longrightarrow H_2O$$

Some bases do not dissociate to yield hydroxide ions directly. For example, ammonia (NH_3) acts as a base by accepting a proton from water, producing an ammonium ion (NH_4^+) and releasing a hydroxide ion.

$$NH_3 + H_2O \longrightarrow NH_4^+ + OH^-$$

A basic solution is one in which the hydrogen ion concentration is lower than the hydroxide ion concentration. Basic solutions turn red litmus paper blue and feel slippery to the touch. In later chapters we will encounter a number of organic bases, such as the purine and pyrimidine bases that are components of nucleic acids.

pH is a convenient measure of acidity

The degree of a solution's acidity is generally expressed in terms of **pH,** defined as the *negative logarithm (base 10) of the hydrogen ion concentration (expressed in moles per liter).*

$$pH = -\log_{10}[H^+]$$

The brackets refer to concentration; therefore the term $[H^+]$ means "the concentration of hydrogen ions," which is expressed in moles per liter because we are interested in the *number* of hydrogen ions per liter. Because the range of possible pH values is very broad, a logarithmic scale (with a tenfold difference between successive units) is more convenient than a linear scale.

Hydrogen ion concentrations are nearly always less than 1 mole per liter. One gram of hydrogen ions dissolved in 1 liter of water (a 1-molar solution) may not sound very impressive, but such a solution would be extremely acidic. The logarithm of a number less than one is a negative number; thus the *negative* logarithm corresponds to a *positive* pH value.

TABLE 2–2 The Relationship Between pH and Hydrogen Ion Concentration

Substance	$[H^+]$*	log $[H^+]$	pH
Gastric juice	0.01, 10^{-2}	−2	2
Pure water, neutral solution	0.0000001, 10^{-7}	−7	7
Household ammonia	0.00000000001, 10^{-11}	−11	11

*$[H^+]$ = hydrogen ion concentration (moles/L)

Whole number pH values are easy to calculate (Table 2–2). For instance, consider our example of pure water, which has a hydrogen ion concentration of 0.0000001 (10^{-7}) moles per liter. The logarithm is −7. The negative logarithm is 7; therefore the pH is 7.

If the hydrogen ion concentration of a solution is known, the hydroxide ion concentration can be easily calculated. The product of the hydrogen ion concentration and the hydroxide ion concentration is 1×10^{-14}:

$$[H^+][OH^-] = 1 \times 10^{-14}$$

In pure (freshly distilled) water, the hydrogen ion concentration is 10^{-7}; therefore the hydroxide concentration is also 10^{-7}. Such a solution, in which the concentrations are equal, is said to be neutral. Acidic solutions (those with an excess of hydrogen ions) have pH values smaller than 7. For example, the hydrogen ion concentration of a solution with pH 1 is ten times that of a solution with pH 2. Basic solutions (those with an excess of hydroxide ions) have pH values greater than 7.

The pH values of some common substances are shown in Figure 2–17. Although some very acidic compartments exist within cells (Chapter 4), most of the interior of an animal or plant cell is neither strongly acidic nor strongly basic but instead an essentially neutral mixture of acidic and basic substances. Although certain bacteria are adapted to life in extremely acidic environments (Chapter 23), a substantial change in pH is incompatible with life for most cells (Fig. 2–18). The pH of most types of plant and animal cells (and their environment) ordinarily ranges from around 7.2 to 7.4.

Buffers minimize pH change

Many homeostatic mechanisms operate to maintain appropriate pH values. For example, the pH of human blood is about 7.4 and must be maintained within very narrow limits. Should the blood become too acidic (for example, as a result of res-

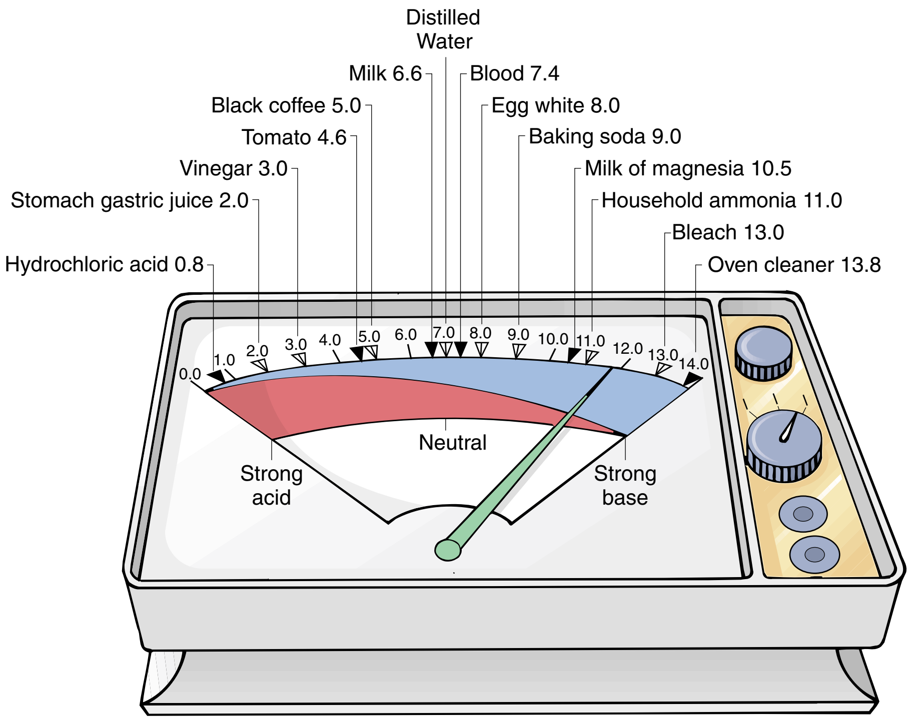


Figure 2–17 **pH values.** The pH meter is an electronic device used to measure the acidity of a solution. A neutral solution (pH 7) has equal concentrations of H^+ and OH^-. Acidic solutions have pH values lower than 7, while basic solutions have pH values higher than 7.

Figure 2–18 **Acid rain damage.** Oxides of sulfur emitted from fossil fuel plants and other industries, and oxides of nitrogen from automobile exhaust, are converted in the moist atmosphere into acids that become dispersed over wide areas. Unlike unpolluted rain (average pH 5.6), the pH of acid rain has been measured at 4.2 and even lower. Plants, such as these trees photographed in the Black Forest, Germany, may be damaged when the resulting increase in soil acidity causes certain minerals, particularly calcium ions, to leach out of the soil. *(Hans Reinhard/Bruce Coleman)*

piratory disease), coma and death may result. Excessive alkalinity can result in overexcitability of the nervous system and even convulsions. Organisms contain many natural buffers. A **buffer** is a substance or combination of substances that resists changes in pH when an acid or base is added. A buffering system includes a weak acid or a weak base (Fig. 2–19). A weak acid or weak base does not ionize completely. That is, at any given instant only a fraction of the molecules are ionized; most are undissociated.

One of the most common buffering systems is found in the blood of vertebrates (see Chapter 44). Carbon dioxide, produced as a waste product of cellular metabolism, enters the blood, the main constituent of which is water. The carbon dioxide reacts with the water to form carbonic acid, a weak acid that dissociates to yield a hydrogen ion and a bicarbonate ion. The buffering system is described by the following expression:

$$\underset{\text{Carbon dioxide}}{CO_2} + \underset{\text{Water}}{H_2O} \rightleftharpoons \underset{\text{Carbonic acid}}{H_2CO_3} \rightleftharpoons H^+ + \underset{\text{Bicarbonate ion}}{HCO_3^-}$$

As indicated by the arrows, all the reactions are reversible. Because carbonic acid is a weak acid, undissociated molecules are always present, as are all the other components of the system. The expression describes the system when it is at equilibrium,

Figure 2–19 **Buffering.** Sodium bicarbonate, which ionizes to form bicarbonate ions (HCO_3^-), is sometimes used as a remedy for excess stomach acid. The bubbles are CO_2 produced by the reaction between a weak acid (citric acid) and the sodium bicarbonate. *(Charles D. Winters)*

TABLE 2–3 Some Biologically Important Ions

Name	Formula	Charge
Sodium	Na^+	1+
Potassium	K^+	1+
Hydrogen	H^+	1+
Magnesium	Mg^{2+}	2+
Calcium	Ca^{2+}	2+
Iron	Fe^{2+} or Fe^{3+}	2+ [iron(II)] or 3+ [iron(III)]
Ammonium	NH_4^+	1+
Chloride	Cl^-	1−
Iodide	I^-	1−
Carbonate	CO_3^{2-}	2−
Bicarbonate	HCO_3^-	1−
Phosphate	PO_4^{3-}	3−
Acetate	CH_3COO^-	1−
Sulfate	SO_4^{2-}	2−
Hydroxide	OH^-	1−
Nitrate	NO_3^-	1−
Nitrite	NO_2^-	1−

when the rates of the forward and reverse reactions are equal and the relative concentrations of the components are not changing. If a system is at equilibrium, it can be "shifted to the right" by adding reactants or removing products. Conversely, it can be "shifted to the left" by adding products or removing reactants. Hydrogen ions are the important products to consider in this system.

The addition of excess hydrogen ions temporarily shifts the system to the left, as they combine with the bicarbonate ions to form carbonic acid. Eventually a new equilibrium is established; at this point the hydrogen ion concentration is similar to the original concentration.

If hydroxide ions are added, they combine with the hydrogen ions to form water, effectively removing a product and thus shifting the system to the right. More carbonic acid then ionizes, replacing the hydrogen ions that were removed.

Organisms contain many weak acids and weak bases, thus maintaining an essential reserve of buffering capacity and avoiding pH extremes.

An acid and a base react to form a salt

When an acid and a base are mixed together, the H^+ of the acid unites with the OH^- of the base to form a molecule of water. The remainder of the acid (an anion) combines with the remainder of the base (a cation) to form a salt. For example, hydrochloric acid reacts with sodium hydroxide to form water and sodium chloride:

$$HCl + NaOH \longrightarrow H_2O + NaCl$$

A **salt** is a compound in which the hydrogen ion of an acid is replaced by some other cation. Sodium chloride, NaCl, is a compound in which the hydrogen ion of HCl has been replaced by the cation Na^+.

When a salt, an acid, or a base is dissolved in water, its dissociated ions can conduct an electrical current; these substances are called **electrolytes.** Sugars, alcohols, and many other substances do not form ions when dissolved in water; they do not conduct an electrical current and are referred to as **nonelectrolytes.**

Cells and extracellular fluids (such as blood) of animals and plants contain a variety of dissolved salts that are the source of the many important mineral ions essential for fluid balance and acid-base balance. The concentrations and relative amounts of the various cations and anions are kept remarkably constant. Any marked change results in impaired cellular functions and may lead to death. Nitrates and ammonium from the soil are the important nitrogen sources for plants. In animals, nerve and muscle function, blood clotting, bone formation, and many other aspects of body function depend on ions. Sodium, potassium, calcium, and magnesium are the chief cations present; chloride, bicarbonate, phosphate, and sulfate are important anions (Table 2–3).

SUMMARY WITH KEY TERMS

I. The chemical composition and metabolic processes of all organisms are very similar. The physical and chemical principles that govern nonliving things also govern organisms.

II. Organisms are made up of small, simple, **inorganic compounds** as well as large, complex, carbon-containing **organic compounds.**

III. An **element** is a substance that cannot be decomposed into simpler substances by normal chemical reactions. Four elements—carbon, hydrogen, oxygen, and nitrogen—make up 96% or more of an organism's mass.

IV. Each **atom** is composed of a nucleus containing **protons** and **neutrons.** In the space outside the nucleus, **electrons** move rapidly in **orbitals** that correspond to energy levels.
   A. An atom is identified by its number of protons **(atomic number).**
   B. Atoms of the same element with different numbers of neutrons (different **atomic masses**) are **isotopes.** Some isotopes are radioactive **(radioisotopes).**
   C. In an electrically neutral atom, the number of protons equals the number of electrons.
   D. The chemical properties of an atom are determined chiefly by the number and arrangement of its most energetic electrons, known as **valence electrons.**

V. Different atoms are joined by chemical bonds to form **compounds.** A **chemical formula** gives the types and relative numbers of atoms in a substance.
   A. One **mole** (the atomic or molecular mass in grams) of any substance contains 6.02×10^{23} atoms, molecules, or ions.
   B. **Covalent bonds** are strong, stable bonds formed when atoms share **valence electrons,** forming **molecules.** A **molecular formula** gives the actual numbers of each type of atom in a molecule; a **structural formula** shows their arrangement.
      1. Covalent bonds are **nonpolar** if the electrons are shared equally between the two atoms.
      2. Covalent bonds are **polar** if one atom is more electronegative (has a greater affinity for electrons) than the other.
   C. An **ionic bond** is formed between a positively charged **cation** and a negatively charged **anion.**
   D. **Hydrogen bonds** are relatively weak bonds formed when a hydrogen atom with a partial positive charge is attracted to an atom (usually oxygen or nitrogen) with a partial negative charge already bonded to another molecule or in another part of the same molecule.

VI. **Oxidation** and **reduction (redox)** reactions are chemical processes in which electrons (and their energy) are transferred from a reducing agent to an oxidizing agent.

VII. Water accounts for a large part of the mass of most organisms. It is important in many chemical reactions within living things and has unique properties that also affect the environment.
   A. Water is a **polar molecule** because one end has a partial positive charge and the other has a partial negative charge.
   B. Because its molecules are polar, water is an excellent **solvent** for ionic or polar **solutes.**
   C. Water molecules are **cohesive** because they form hydrogen bonds with each other; they are also **adhesive** through hydrogen bonding to substances with ionic or polar regions.
   D. Water has a high **specific heat,** which helps organisms maintain a relatively constant internal temperature; this property also helps keep the ocean and other large bodies of water at a constant temperature.
   E. Water has a high **heat of vaporization.** Molecules entering the vapor phase carry a great deal of heat, which accounts for **evaporative cooling.**
   F. The fact that ice is less dense than liquid water makes the environment less extreme.
   G. Water has a slight tendency to **ionize,** that is, to dissociate to form hydrogen ions (protons, H^+) and hydroxide ions (OH^-).

VIII. **Acids** are proton (H^+) donors; **bases** are proton acceptors. Many bases dissociate in solution to yield hydroxide ions, which then accept protons.
   A. The **pH scale** is a logarithmic expression of the hydrogen ion concentration of a solution. As a solution becomes more acidic, its pH falls below 7 (neutrality). As a solution becomes more basic (alkaline), its pH rises above 7.
   B. A buffering system is based on a weak acid or a weak base. A **buffer** resists changes in the pH of a solution when acids or bases are added.
   C. A **salt** is a compound in which the hydrogen atom of an acid is replaced by some other cation. Salts provide the many mineral ions essential for life functions.

POST-TEST

1. Which of the following elements is mismatched with its properties or function? (a) carbon—forms the backbone of organic compounds (b) nitrogen—component of proteins (c) hydrogen—very electronegative (d) oxygen—can participate in hydrogen bonding (e) all of the above are correctly matched
2. Which of the following applies to a neutron? (a) positive charge and located in an orbital (b) negligible mass and located in the nucleus (c) positive charge and located in the nucleus (d) uncharged and located in the nucleus (e) uncharged and located in an orbital
3. $^{32}_{15}P$, a radioactive form of phosphorus, has (a) an atomic number of 32 (b) an atomic mass of 15 (c) an atomic mass of 47 (d) 32 electrons (e) 17 neutrons
4. Which of the following facts allows you to determine that atom A and atom B are isotopes of the same element? (a) they each have six protons (b) they each have four neutrons (c) in each, the sum of their electrons and neutrons is 14 (d) they each have four valence electrons (e) they each have a mass number of 14
5. Sodium and potassium behave similarly in chemical reactions. This is because (a) they have the same number of neutrons (b) each has a single valence electron (c) they have the same atomic mass (d) they have the same number of electrons (e) they have the same number of protons
6. The orbitals comprising an atom's valence electron shell (a) are arranged as concentric spheres (b) contain the atom's least energetic electrons (c) may change shape when covalent bonds are formed (d) never contain more than one electron each (e) more than one of the above is correct
7. Which of the following bonds and properties are correctly matched? (a) ionic bonds—strong only if the participating ions are hydrated (b) hydrogen bonds—responsible for bonding oxygen and hydrogen to form a single water molecule (c) polar covalent bonds—can occur between two atoms of the same element (d) covalent bonds—may be single, double, or triple (e) hydrogen bonds—stronger than covalent bonds
8. In a redox reaction (a) energy is transferred from a reducing agent to an oxidizing agent (b) a reducing agent becomes oxidized as it accepts an electron (c) an oxidizing agent accepts a proton (d) a reducing agent donates a proton (e) the electrons in an atom move from its valence shell to a shell closer to its nucleus
9. A solution with a pH of 2 has a hydrogen ion concentration that is ____________ the hydrogen ion concentration of a solution with a pH of 4. (a) 1/2 (b)1/100 (c) two times (d) ten times (e) one hundred times
10. The high heat of vaporization of water accounts for (a) evaporative cooling (b) the fact that ice floats (c) the fact that heat is liberated when ice forms (d) the cohesive properties of water (e) capillary action

11. NaOH and HCl react to form Na^+, Cl^-, and water. Which of the following statements is true? (a) Na^+ is an anion, and Cl^- is a cation (b) Na^+ and Cl^- are both anions (c) a hydrogen bond can form between Na^+ and Cl^- (d) Na^+ and Cl^- are electrolytes (e) Na^+ is an acid, and Cl^- is a base

REVIEW QUESTIONS

1. What is the relationship between molecules and compounds? Are all compounds composed of molecules?
2. What are the ways an atom or molecule can become an anion or a cation?
3. What is a radioisotope? Why is it able to substitute for an ordinary (nonradioactive) atom of the same element in a molecule? What are some of the ways radioisotopes are used in biological research?
4. How do ionic and covalent bonds differ?
5. Why does water form hydrogen bonds? List some of the properties of water that result from hydrogen bonding. How do these properties contribute to the role of water as an essential component of organisms?
6. How can weak forces, such as hydrogen bonds, have significant effects in organisms?
7. A solution has a hydrogen ion concentration of 0.01 moles/liter. What is its pH? What is its hydroxide ion concentration? How would this solution differ from one with a pH of 1?
8. Why are buffers important in organisms? Give a specific example of how a buffering system works.
9. Differentiate clearly among acids, bases, and salts.
10. Why must oxidation and reduction occur simultaneously? Why are redox reactions important in some energy transfers?
11. Describe a reversible reaction that is at equilibrium. What would be the consequences of adding or removing a reactant or a product?

YOU MAKE THE CONNECTION

1. Element A has two electrons in its valence shell (which is complete when it contains eight electrons). Would you expect element A to share, donate, or accept electrons? What would you expect of element B, which has four valence electrons, and element C, which has seven?
2. A hydrogen bond formed between two water molecules is only about one-twentieth as strong as a covalent bond between hydrogen and oxygen. In what ways would the physical properties of water be different if these hydrogen bonds were stronger (e.g., one-tenth the strength of covalent bonds)?
3. Consider the following reaction (in water).

$$HCl \longrightarrow H^+ + Cl^-$$

Name the reactant(s) and product(s). Does the expression indicate that the reaction is reversible? Could HCl be used as a buffer?

RECOMMENDED READINGS

Atkins, P.W. *Periodic Kingdom.* Basic Books, Harper Collins Publishers, New York, 1995. In this imaginative work the periodic table is described as a landscape inhabited by elements whose properties are determined by the region in which they reside.

Hedin, L.O. and G.E. Likens. "Atmosphere Dust and Acid Rain." *Scientific American,* Vol. 275, No. 6 Dec. 1996. This article examines the idea that recent reductions in basic compounds attached to dust particles in the atmosphere may be increasing the environmental damage done by acid rain.

Joesten, M.D. and J.L. Wood. *World of Chemistry.* Saunders College Publishing, Philadelphia, 1996. A very readable introduction to chemistry.

Zimmer, C. "Wet, Wild and Weird." *Discover,* Dec. 1992. Computer simulations illustrate the many ways water molecules can interact through hydrogen bonding.

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