Life as we know it would be impossible without water, a small inorganic compound with unique properties.

Nigel Cattlin/Holt Studios International/Science Photo Library/Photo Researchers, Inc.

Study Plan

2.1 The Organization of Matter: Elements and Atoms

Living organisms are composed of about 25 key elements

Elements are composed of atoms, which combine to form molecules

2.2 Atomic Structure

The atomic nucleus contains protons and neutrons

The nuclei of some atoms are unstable and tend to break down to form simpler atoms

The electrons of an atom occupy orbitals around the nucleus

Orbitals occur in discrete layers around an atomic nucleus

The number of electrons in the outermost energy level of an atom determines its chemical activity

2.3 Chemical Bonds

Ionic bonds are multidirectional and vary in strength

Covalent bonds are formed by electrons in shared orbitals

Unequal electron sharing results in polarity

Polar molecules tend to associate with each other and exclude nonpolar molecules

Hydrogen bonds also involve unequal electron sharing

Van der Waals forces are weak attractions over very short distances

Bonds form and break in chemical reactions

2.4 Hydrogen Bonds and the Properties of Water

A lattice of hydrogen bonds gives water unusual properties

The hydrogen-bond lattice of water contributes to polar and nonpolar environments in and around cells

The small size and polarity of its molecules makes water a good solvent

The hydrogen-bond lattice gives water other life-sustaining properties as well

2.5 Water Ionization and Acids, Bases, and Buffers

Substances act as acids or bases by altering the concentrations of H^+ and OH^- ions in water

Buffers help keep pH under control

2 Life, Chemistry, and Water

Why It Matters

We—like all plants, animals, and other organisms—are collections of atoms and molecules linked together by chemical bonds. Our chemical nature makes it impossible to understand biology without knowledge of basic chemistry and chemical interactions.

For example, the element selenium is a natural ingredient of rocks and soil. In minute amounts it is necessary for the normal growth and survival of humans and many other animals, but high concentrations of selenium are toxic. In 1983, thousands of dead or deformed waterfowl were discovered at the Kesterson Wildlife Refuge in the San Joaquin Valley of California. The deaths and deformities were traced to high concentrations of selenium in the environment, alerting the public and scientists alike to the dangers of this element to all animals. Decades of irrigation had washed selenium-containing chemicals from the soil into the water of the refuge. With the problem identified, engineers have diverted agricultural drainage water from the area, and the Kesterson refuge is now being restored.

The study of selenium and its biological effects has suggested a possible way to prevent it from accumulating in the environment. In 1996, Norman Terry and his coworkers at the University of California at Berkeley started a large-scale experiment designed to test natural methods for removing excess selenium from contaminated soils. Terry found that some wetland plants could remove up to 90% of the selenium in wastewater from a gasoline refinery. The plants convert much of the selenium into a relatively nontoxic gas, methyl selenide, which can pass into the atmosphere without harming plants and animals.

To test further the ability of plants to remove selenium, Terry and his coworkers grew wetland plants in 10 experimental plots watered by runoff from agricultural irrigation **(Figure 2.1).** The researchers measured how much selenium remained in the soil of the plots, how much was incorporated into plant tissues, and how much escaped into the air as a gas. Terry's results indicate that before the runoff trickles through to local ponds, the plants in his plots reduce selenium to nontoxic levels, less than 2 parts per billion. Such applications of chemical and biological knowledge to decontaminate polluted environments are known as **bioremediation.** They could help safeguard our food supplies, our health, and the environment.

The selenium example shows the importance of understanding and applying chemistry in biology. However, reactions involving selenium are only a few of the many thousands of chemical reactions that take place inside living organisms. Decades of research have taught us much about these reactions and have confirmed that the same laws of chemistry and physics govern both living and nonliving things. We can therefore apply with confidence information from chemical experiments in the laboratory to the processes inside living organisms. An understanding of the relationship between the structure of chemical substances and their behavior is the first step toward learning biology, and this knowledge will help you appreciate the benefits and risks of applying chemistry to human affairs.

Photo by Gary Head for Norman Terry, University of California, Berkeley

Figure 2.1
Researcher Norman Terry in an experimental wetlands plot in Corcoran, California.
Terry is testing the ability of cattails, bulrushes, and marsh grasses to reduce selenium contamination in water draining from irrigated fields.

2.1 The Organization of Matter: Elements and Atoms

Selenium is an example of an **element**—a pure substance that cannot be broken down into simpler substances by ordinary chemical or physical techniques. All **matter** of the universe—anything that occupies space and has mass—is composed of elements and combinations of elements. Ninety-two different elements occur naturally on Earth, and more than fifteen artificial elements have been synthesized in the laboratory.

Living Organisms Are Composed of about 25 Key Elements

Four elements—carbon, hydrogen, oxygen, and nitrogen—make up more than 96% of the weight of living organisms. Seven other elements—calcium, phosphorus, potassium, sulfur, sodium, chlorine, and magnesium—contribute most of the remaining 4%. Several other elements occur in organisms in quantities so small ($<$0.01%) that they are known as **trace elements. Figure 2.2** compares the relative proportions of different elements in a human, a plant, Earth's crust, and seawater, and lists the most important trace elements in a human. The proportions of elements in living organisms, as represented by the human and the plant, differ markedly from those of Earth's crust and seawater; these differences reflect the highly ordered chemical structure of living organisms.

Trace elements are vital to normal biological functions. For example, iodine makes up only about 0.0004% of a human's weight. However, a lack of iodine in the human diet severely impairs the function of the thyroid gland, which produces hormones that regulate metabolism and growth. Symptoms of iodine deficiency include lethargy, apathy, and sensitivity to cold temperatures. Prolonged iodine deficiency causes a *goiter,* a condition in which the thyroid gland enlarges so much that the front of the neck swells grotesquely. Once a common condition, goiter has almost been eliminated by adding iodine to table salt, especially in regions where soils are iodine-deficient.

Elements Are Composed of Atoms, Which Combine to Form Molecules

Elements are composed of individual **atoms**—the smallest units that retain the chemical and physical properties of an element. Any given element has only one type of atom. Several million atoms arranged side by side would be needed to equal the width of the period at the end of this sentence.

Atoms are identified by a standard one- or two-letter symbol. The element carbon is identified by the single letter *C*, which stands for both the carbon atom and the element; iron is identified by the two-letter symbol *Fe* (*ferrum* = iron). **Table 2.1** lists the chemical symbols of these and other atoms common in living organisms.

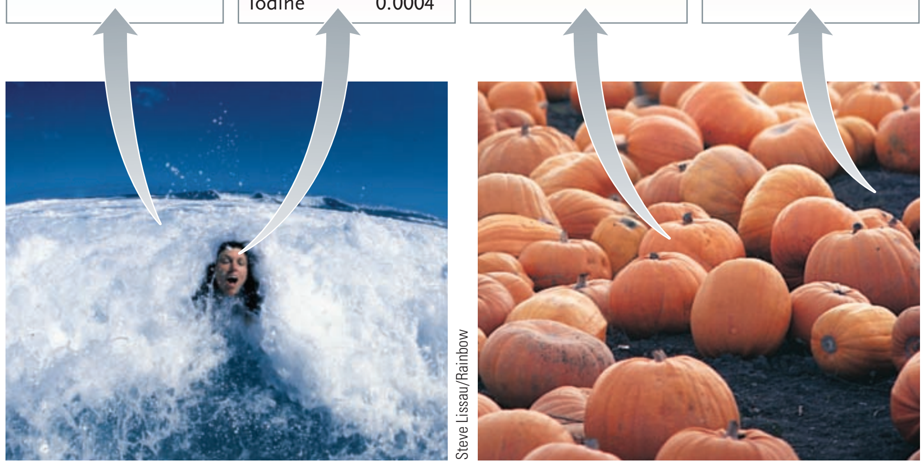
Atoms combined chemically in fixed numbers and ratios form the **molecules** of living and nonliving matter. For example, the oxygen we breathe is a molecule formed from the chemical combination of two oxygen atoms; a molecule of the carbon dioxide that we exhale contains one carbon atom and two oxygen atoms. The name of a molecule is written in chemical shorthand as a **formula,** using the standard symbols for the elements and using subscripts to indicate the number of atoms of each element in the molecule. The subscript is omitted for atoms that occur only once in a molecule. For example, the formula for an oxygen molecule is written as O_2; for a carbon dioxide molecule the formula is CO_2.

Seawater	
Oxygen	88.3
Hydrogen	11.0
Chlorine	1.9
Sodium	1.1
Magnesium	0.1
Sulfur	0.09
Potassium	0.04
Calcium	0.04
Carbon	0.003
Silicon	0.0029
Nitrogen	0.0015
Strontium	0.0008

Human	
Oxygen	65.0
Carbon	18.5
Hydrogen	9.5
Nitrogen	3.3
Calcium	2.0
Phosphorus	1.1
Potassium	0.35
Sulfur	0.25
Sodium	0.15
Chlorine	0.15
Magnesium	0.05
Iron	0.004
Iodine	0.0004

Pumpkin	
Oxygen	85.0
Hydrogen	10.7
Carbon	3.3
Potassium	0.34
Nitrogen	0.16
Phosphorus	0.05
Calcium	0.02
Magnesium	0.01
Iron	0.008
Sodium	0.001
Zinc	0.0002
Copper	0.0001

Earth's crust	
Oxygen	46.6
Silicon	27.7
Aluminum	8.1
Iron	5.0
Calcium	3.6
Sodium	2.8
Potassium	2.6
Magnesium	2.1
Other elements	1.5



Steve Lissau/Rainbow

Jack Carey

Figure 2.2 **The proportions by mass of different elements in seawater, the human body, a fruit, and Earth's crust.** Trace elements in humans include boron, chromium, cobalt, copper, fluorine, iodine, iron, manganese, molybdenum, selenium, tin, vanadium, and zinc, as well as variable traces of other elements.

Table 2.1 **Atomic Number and Mass Number of the Most Common Elements in Living Organisms**

Element	Symbol	Atomic Number	Mass Number of the Most Common Form
Hydrogen	H	1	1
Carbon	C	6	12
Nitrogen	N	7	14
Oxygen	O	8	16
Sodium	Na	11	23
Magnesium	Mg	12	24
Phosphorus	P	15	31
Sulfur	S	16	32
Chlorine	Cl	17	35
Potassium	K	19	39
Calcium	Ca	20	40
Iron	Fe	26	56
Iodine	I	53	127

Molecules whose component atoms are different (such as carbon dioxide) are called **compounds.** The chemical and physical properties of compounds are typically distinct from those of their atoms or elements. For example, we all know that water is a liquid at room temperature. We also know that water does not burn. However, the properties of the individual elements of water—hydrogen and oxygen—are quite different. Hydrogen and oxygen are gases at room temperature, and both are highly reactive.

Study Break

Distinguish between an element and an atom and between a molecule and a compound.

2.2 Atomic Structure

Each element consists of one type of atom. However, all atoms share the same basic structure **(Figure 2.3).** Each atom consists of an **atomic nucleus**, surrounded by one or more smaller, fast-moving particles called **electrons.** Although the electrons occupy more than 99.99% of the space of an atom, the nucleus makes up more than 99.99% of its total mass.

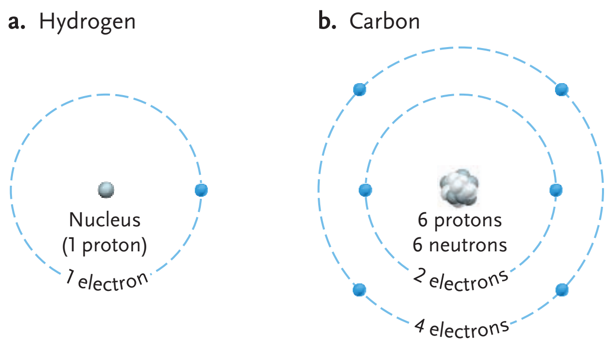


Figure 2.3
Atomic structure. The nucleus of an atom contains one or more protons and, except for the most common form of hydrogen, a similar number of neutrons. Fast-moving electrons, in numbers equal to the protons, surround the nucleus. **(a)** The most common form of hydrogen, the simplest atom, has a single proton in its nucleus and a single electron. **(b)** Carbon, a more complex atom, has a nucleus surrounded by electrons at two levels. The electrons in the outer level follow more complex pathways than shown here.

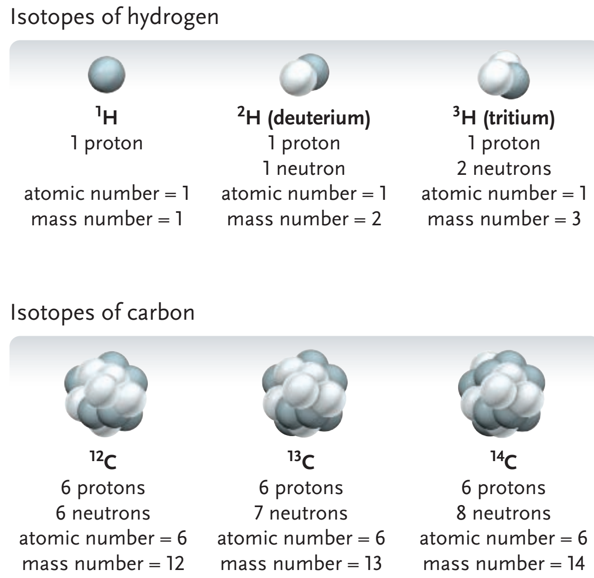


Figure 2.4
The atomic nuclei of hydrogen and carbon isotopes. Note that isotopes of an atom have the same atomic number but different mass numbers.

The Atomic Nucleus Contains Protons and Neutrons

All atomic nuclei contain one or more positively charged particles called **protons.** The number of protons in the nucleus of each kind of atom is referred to as the **atomic number.** This number does not vary and thus specifically identifies the atom. The smallest atom, hydrogen, has a single proton in its nucleus, so its atomic number is 1. The heaviest naturally occurring atom, uranium, has 92 protons in its nucleus and therefore has an atomic number of 92. Similarly, carbon with six protons, nitrogen with seven protons, and oxygen with eight protons have atomic numbers of 6, 7, and 8, respectively (see Table 2.1).

With one exception, the nuclei of all atoms also contain uncharged particles called **neutrons,** which occur in variable numbers approximately equal to the number of protons. The single exception is the most common form of hydrogen, which has a nucleus that contains only a single proton. There are two less common forms of hydrogen as well. One form, named deuterium, has a neutron in its nucleus in addition to a single proton. The other form, named tritium, has two neutrons and a single proton.

Other atoms also have common and less common forms with different numbers of neutrons. For example, the most common form of the carbon atom has six protons and six neutrons in its nucleus, but about 1% of carbon atoms have six protons and seven neutrons in their nuclei and an even smaller percentage has six protons and eight neutrons.

These distinct forms of the atoms of an element, all with the same number of protons but different numbers of neutrons, are called **isotopes (Figure 2.4).** The various isotopes of an atom differ in mass and other physical characteristics, but all have essentially the same chemical properties. Therefore, organisms can use any hydrogen or carbon isotope, for example, without a change in their chemical reactions.

A neutron and a proton have almost the same mass, about 1.66×10^{-24} grams (g). This mass is defined as a standard unit, the **dalton,** named after John Dalton, a nineteenth-century English scientist who contributed to the development of atomic theory. Atoms are assigned a **mass number** based on the total number of protons and neutrons in the atomic nucleus (see Table 2.1). Electrons are ignored in determinations of atomic mass because the mass of an electron, at only 1/1800 of the mass of a proton or neutron, does not contribute significantly to the mass of an atom. Thus, the mass number of the hydrogen isotope with one proton in its nucleus is 1, and its mass is 1 dalton. The mass number of the hydrogen isotope deuterium is 2, and the mass number of tritium is 3. The carbon isotope with six protons and six neutrons in its nucleus has a mass number of 12; the isotope with six protons and seven neutrons has a mass number of 13, and the isotope with six protons and eight neutrons has a mass number of 14 (see Figure 2.4). These carbon mass numbers are written as ^{12}C, ^{13}C, and ^{14}C, or carbon-12, carbon-13, and carbon-14, respectively. However, all the carbon isotopes have the same atomic number of 6, because this number reflects only the number of protons in the nucleus.

You might wonder about the meaning of mass as compared to weight. Mass is the amount of matter in an object, whereas weight measures the pull of gravity on an object. Mass is constant, but the weight of an object may vary because of differences in gravity. For example, the mass of a piece of lead is the same on Earth and in outer space, but the same piece of lead that weighs 1 kilogram (kg) on Earth is weightless in an orbiting spacecraft, even though its mass remains the same. However, as long as an object is on Earth's surface, its mass and weight are equivalent. Thus, we can weigh an object in the laboratory and be assured that its weight accurately reflects its mass.

The Nuclei of Some Atoms Are Unstable and Tend to Break Down to Form Simpler Atoms

The nuclei of some isotopes are unstable and break down, or *decay,* giving off particles of matter and energy that can be detected as **radioactivity.** The decay transforms the unstable, radioactive isotope—called a **radioisotope**—into an atom of another element. The decay continues at a steady, clocklike rate, with a constant proportion of the radioisotope breaking down at any instant. The rate of decay is not affected by chemical reactions or environmental conditions such as temperature or pressure. For example, the carbon isotope ^{14}C is unstable and undergoes radioactive decay in which one of its neutrons splits into a proton and an electron. The electron is ejected from the nucleus, but the proton is retained, giving a new total of seven protons and seven neutrons, which is characteristic of the most common form of nitrogen. Thus, the decay transforms the carbon atom into an atom of nitrogen.

Because unstable isotopes decay at a clocklike rate, they can be used to estimate the age of organic material, rocks, or fossils that contain them. These techniques have been vital in dating animal remains and tracing evolutionary lineages, as described in Chapter 22. Isotopes are also used in biological research as **tracers** to label molecules so that they can be tracked as they pass through biochemical reactions. Radioactive isotopes of carbon (^{14}C), phosphorus (^{32}P), and sulfur (^{35}S) can be traced easily by their radioactivity. A number of stable, nonradioactive isotopes, such as ^{15}N (called heavy nitrogen), can be detected by their mass differences and have also proved valuable as tracers in biological experiments. *Focus on Research* describes some applications of radioisotopes in research and medicine.

The Electrons of an Atom Occupy Orbitals around the Nucleus

In an atom, the number of electrons surrounding the nucleus is equal to the number of protons in the nucleus. An electron carries a negative charge that is exactly equal and opposite to the positive charge of a proton, balancing the positive and negative charges and making the total structure of an atom electrically neutral.

An atom is often drawn in a simple way with electrons orbiting the nucleus similar to planets orbiting a sun. The reality is different. Electrons are in constant motion around the nucleus, moving at speeds that approach the speed of light. At any instant, an electron may be in any location with respect to its nucleus, from the immediate vicinity of the nucleus to practically infinite space. An electron moves so fast that it almost occupies all the locations at the same time; however, it passes through some locations much more frequently than others. The locations where an electron occurs most frequently around the atomic nucleus define a path called an orbital. An **orbital** is essentially the region of space where the electron "lives" most of the time. Although either one or two electrons may occupy an orbital, the most stable and balanced condition occurs when an orbital contains a pair of electrons.

Electrons are maintained in their orbitals by a combination of attraction to the positively charged nucleus and mutual repulsion because of their negative charge. The orbitals take different shapes depending on their distance from the nucleus and their degree of repulsion by electrons in other orbitals.

Under certain conditions, electrons may pass from one orbital to another within an atom, enter orbitals shared by two or more atoms, or pass completely from orbitals in one atom to orbitals in another. As discussed later in this chapter, the ability of electrons to move from one orbital to another underlies the chemical reactions that combine atoms into molecules.

Orbitals Occur in Discrete Layers around an Atomic Nucleus

Within an atom, electrons are found in regions of space called **energy levels,** or more simply, **shells.** Within each energy level, electrons are grouped into orbitals. The lowest energy level of an atom, the one nearest the nucleus, may be occupied by a maximum of two electrons in a single orbital **(Figure 2.5a).** This orbital, which has a spherical shape, is called the 1*s* orbital. (The "1" signifies that the orbital is in the energy level closest to the nucleus, and the "*s*" signifies the shape of the orbital, in this case, spherically symmetric around the nucleus.) Hydrogen has one electron in this orbital, and helium has two.

Atoms with atomic numbers between 3 (lithium) and 10 (neon) have two energy levels, with two electrons in the 1*s* orbital and one to eight electrons in orbitals at the next highest energy level. The electrons at this second energy level occupy one spherical orbital, called the 2*s* orbital **(Figure 2.5b),** and as many as three orbitals that are pushed into a dumbbell shape by repulsions between electrons, called 2*p* orbitals **(Figure 2.5c). Figure 2.5d** shows the orbitals for neon.

Larger atoms have more energy levels. The third energy level, which may contain as many as 18 electrons in 9 orbitals, includes the atoms from sodium (11 electrons) to argon (18 electrons). (**Figure 2.6** shows the 18 elements that have electrons in the lowest three energy levels only.) The fourth energy level may contain as many as 32 electrons in 16 orbitals. In all cases, the total number of electrons in the orbitals is matched by the number of protons in the nucleus. However, no matter what the size of an atom, the outermost energy level typically contains one to eight electrons occupying a maximum of four orbitals.

Focus on Research

Basic Research: Using Radioisotopes to Trace Reactions and Save Lives

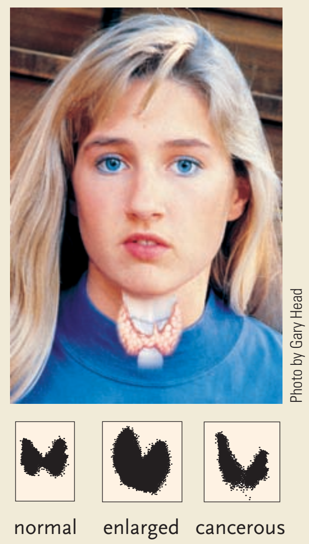
In 1896, the French physicist Henri Becquerel wrapped a rock containing uranium in paper and tucked it into a desk drawer on top of a case containing an unexposed photographic plate. When he opened the case containing the plate a few days later, he was surprised to find an image of the rock on the plate—apparently caused by energy emitted from the rock. One of his coworkers, Marie Curie, named the phenomenon "radioactivity." Although radioactivity can be dangerous to life (more than one researcher, including Marie Curie, has died from its effects), it has been harnessed and put to highly productive use for scientific and medical purposes.

The radiation released by unstable isotopes can be detected by placing a photographic film over samples containing the isotopes (as Becquerel discovered) or by using an instrument known as a *scintillation counter.* These techniques allow researchers to use isotopes as tracers in chemical reactions. Typically, organisms are exposed to a reactant chemical that has been "labeled" with a radioactive isotope such as ^{14}C or ^{3}H. After being exposed to the tracer, the chemical products in which the isotope appears, and their sequence of appearance, can be detected by their radioactivity and identified.

For example, algae and plants use carbon dioxide (CO_2) as a raw material in photosynthesis. To trace the reactions of photosynthesis, Melvin Calvin and his coworkers grew algal cells in a medium with CO_2 that contained the radioisotope ^{14}C. Then they extracted various substances from the cells at intervals, separated them on a piece of paper based on their different solubilities in particular solvents, and placed the paper on a photographic film. The particular substances that exposed spots on the film because they were radioactive, as well as their order of appearance in the cells, allowed the researchers to piece together the sequence of reactions in photosynthesis, as described and illustrated in the *Focus on Research* in Chapter 9.

Radioisotopes are widely used in medicine to diagnose and cure disease, to produce images of diseased body organs, and, as in biological research, to trace the locations and routes followed by individual substances marked for identification by radioactivity. One example of their use in diagnosis is in the evaluation of thyroid gland disease. The thyroid is the only structure in the body that absorbs iodine in quantity. The size and shape of the thyroid, which reflect its health, are measured by injecting a small amount of a radioactive iodine isotope into the patient's bloodstream. After the isotope is concentrated in the thyroid, the gland is then scanned by an apparatus that uses the radioactivity to produce an image of the gland on a photographic film. Examples of what the scans may show are presented in the figure. Another application uses the fact that radioactive thallium is not taken up by regions of the heart muscle with poor circulation to detect coronary artery disease. Other isotopes are used to detect bone injuries and defects, including injured, arthritic, or abnormally growing segments of bone.

Treatment of disease with radioisotopes takes advantage of the fact that radioactivity in large doses can kill cells (radiation generates highly reactive chemical groups that break and disrupt biological molecules). Dangerously overactive thyroid glands are treated by giving patients a dose of radioactive iodine calculated to destroy just enough thyroid cells to reduce activity of the gland to normal levels. In radiation therapy, cancer cells are killed by bombarding them with radiation emitted by radium-226 or cobalt-60. As much as is possible, the radiation is focused on the tumor to avoid destroying nearby healthy tissues. In some forms of chemotherapy for cancer, patients are given radioactive substances at levels that kill cancer cells without also killing the patient.



Scans of human thyroid glands after iodine-123 was injected into the bloodstream. The radioactive iodine becomes concentrated in the thyroid gland.

The Number of Electrons in the Outermost Energy Level of an Atom Determines Its Chemical Activity

The electrons in an atom's outermost energy level are known as **valence electrons** (*valentia* = power or capacity). Atoms in which the outermost energy level is not completely filled with electrons tend to be chemically reactive; those with a completely filled outermost energy level are nonreactive, or inert. For example, hydrogen has a single, unpaired electron in its outermost and only energy level, and it is highly reactive; helium has two valence electrons filling its single orbital, and it is inert. For atoms with two or more energy levels, only those with unfilled outer energy levels are reactive. Those with eight electrons completely fill-

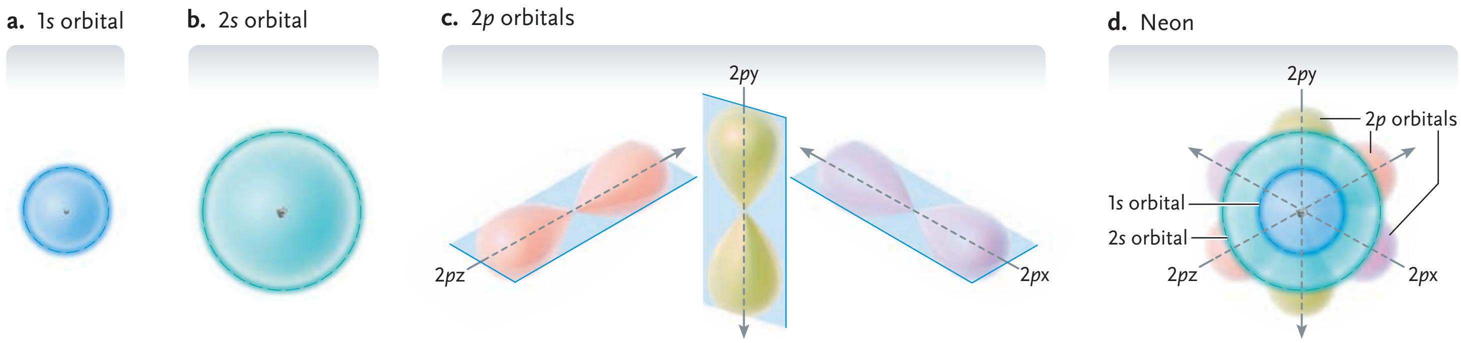


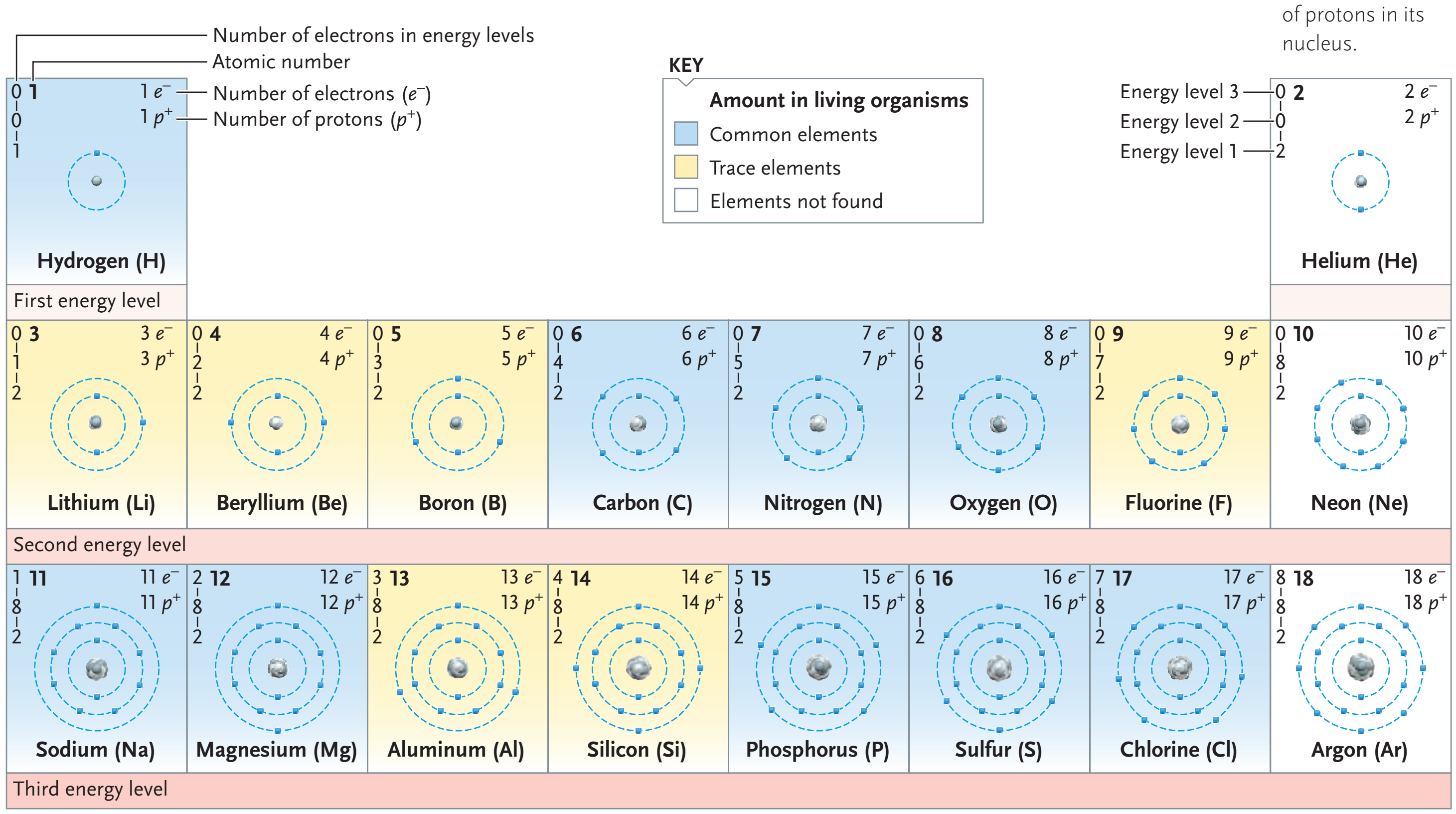
Figure 2.5
Electron orbitals. (a) The single 1*s* orbital of hydrogen and helium approximates a sphere centered on the nucleus. **(b)** The 2*s* orbital. **(c)** The 2*p* orbitals lie in the three planes x, y, and z, each at right angles to the others. **(d)** In atoms with two energy levels, such as neon, the lowest energy level is occupied by a single 1*s* orbital as in hydrogen and helium. The second, higher energy level is occupied by a maximum of four orbitals—a spherical 2*s* orbital and three dumbbell-shaped 2*p* orbitals.

ing the four orbitals of the outer energy level, such as neon and argon, are stable and chemically unreactive (see Figure 2.6).

Atoms with outer energy levels that contain electrons near the stable numbers tend to gain or lose electrons to reach the stable configuration. For example, sodium has two electrons in its first energy level, eight in the second, and one in the third and outermost level (see Figure 2.6). The outermost electron is readily lost to another atom, giving the sodium atom a stable second energy level (now the outermost level) with eight electrons. Chlorine, with seven electrons in its outermost energy level, tends to take up an electron from another atom to attain the stable number of eight electrons.

Atoms that differ from the stable configuration by more than one or two electrons tend to attain stability by *sharing* electrons in joint orbitals with other atoms rather than by gaining or losing electrons completely. Among the atoms that form biological molecules, electron sharing is most characteristic of carbon, which has four electrons in its outer energy level and thus falls at the midpoint between the tendency to gain or lose electrons. Oxygen, with six electrons in its outer level, and nitrogen, with five electrons in its outer level, also share electrons readily. Hydrogen may either share or lose its single electron. The relative tendency to gain, share, or lose valence electrons underlies the chemical bonds and forces that hold the atoms of molecules together.

Figure 2.6
The atoms with electrons distributed in one, two, or three energy levels. The atomic number of each element (shown in boldface in each panel) is equivalent to the number of protons in its nucleus.



Study Break

1. Where are protons, electrons, and neutrons found in an atom?
2. The isotopes carbon-11 and oxygen-15 do not occur in nature, but they can be made in the laboratory. Both are used in a medical imaging procedure called positron emission tomography. Give the number of protons and neutrons in carbon-11 and in oxygen-15.
3. What determines the chemical reactivity of an atom?

2.3 Chemical Bonds

Atoms of inert elements, such as helium, neon, and argon, occur naturally in uncombined forms, but atoms of reactive elements tend to combine into molecules by forming **chemical bonds.** The four chemical linkages that are important in biological molecules are **ionic bonds,** resulting from electrical attractions between atoms that have lost or gained electrons; **covalent bonds,** formed by electron sharing between atoms; **hydrogen bonds,** noncovalent bonds formed by unequal electron sharing between hydrogen atoms and oxygen, nitrogen, or sulfur atoms; and **van der Waals forces,** weak molecular attractions over short distances.

Ionic Bonds Are Multidirectional and Vary in Strength

Ionic bonds form between atoms that gain or lose valence electrons completely. A sodium atom (Na) readily loses a single electron to achieve a stable outer energy level, and chlorine (Cl) readily gains an electron:

$$\mathrm{Na}\cdot + \cdot\underset{\cdot\cdot}{\overset{\cdot\cdot}{\mathrm{Cl}}}: \rightarrow \mathrm{Na}^{+}\ :\underset{\cdot\cdot}{\overset{\cdot\cdot}{\mathrm{Cl}}}:^{-}$$

(The dots in the preceding formula represent the electrons in the outermost energy level.) After the transfer, the sodium atom, now with 11 protons and 10 electrons, carries a single positive charge. The chlorine atom, now with 17 protons and 18 electrons, carries a single negative charge. In this charged condition, the sodium and chlorine atoms are called **ions** instead of atoms and are written as Na^+ and Cl^- **(Figure 2.7).** A positively charged ion such as Na^+ is

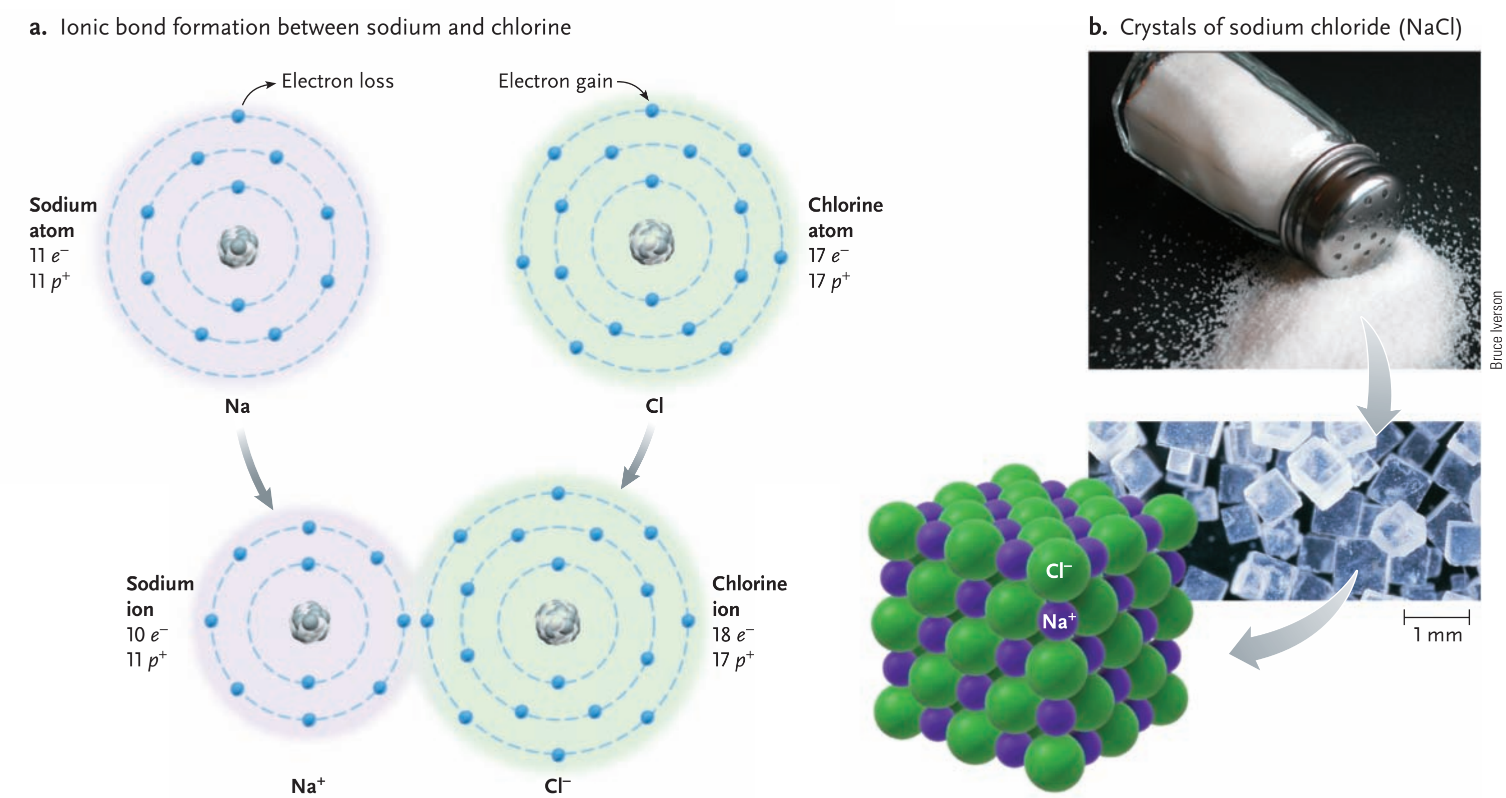


Figure 2.7
Formation of an ionic bond. **(a)** Sodium, with one electron in its outermost energy level, readily loses that electron to attain a stable state in which its second energy level, with eight electrons, becomes the outer level. Chlorine, with seven electrons in its outer energy level, readily gains an electron to attain the stable number of eight. The transfer creates the ions Na^+ and Cl^-. **(b)** The combination forms sodium chloride (NaCl), common table salt.

called a **cation**, and a negatively charged ion such as Cl^- is called an **anion.** The difference in charge between cations and anions creates an attraction—the ionic bond—that holds the ions together in solid NaCl (sodium chloride).

Many other atoms that differ from stable outer energy levels by one electron, including hydrogen, can gain or lose electrons completely to form ions and ionic bonds. When a hydrogen atom loses its single electron to form a hydrogen ion (H^+), it consists of only a proton and is often simply called a proton to reflect this fact. A number of atoms with outer energy levels that differ from the stable number by two or three electrons, particularly metallic atoms such as calcium (Ca^{2+}), magnesium (Mg^{2+}), and iron (Fe^{2+} or Fe^{3+}), also lose their electrons readily to form cations and to join in ionic bonds with anions.

Ionic bonds are common among the forces that hold ions, atoms, and molecules together in living organisms because these bonds have three key features: (1) they exert an attractive force over greater distances than any other chemical bond, (2) their attractive force extends in all directions, and (3) they vary in strength depending on the presence of other charged substances. That is, in some systems, ionic bonds form in locations that exclude other charged substances, setting up strong and stable attractions that are not easily disturbed. For example, iron ions are stabilized by ionic bonds in the interior of the large biological molecule hemoglobin, where the ions are key to the distinctive chemical properties of that molecule. In other systems, particularly at molecular surfaces exposed to water molecules, ionic bonds are relatively weak, allowing ionic attractions to be established or broken quickly. For example, as part of their activity in speeding biological reactions, many enzymatic proteins bind and release molecules by forming and breaking relatively weak ionic bonds.

Covalent Bonds Are Formed by Electrons in Shared Orbitals

Covalent bonds form when atoms share a pair of valence electrons rather than gaining or losing them. The formation of molecular hydrogen, H_2, by two hydrogen atoms is the simplest example of the sharing mechanism. If two hydrogen atoms collide, the single electron of each atom may join in a new, combined two-electron orbital that surrounds both nuclei. The two electrons fill the orbital; thus, the hydrogen atoms tend to remain linked stably together. The linkage formed by the shared orbital is a covalent bond.

In molecular diagrams, a covalent bond is designated by a pair of dots or a single line that represents a pair of shared electrons. For example, in H_2, the covalent bond that holds the molecule together is represented as H:H or H—H.

Unlike ionic bonds, which extend their attractive force in all directions, the shared orbitals that form covalent bonds extend between atoms at discrete angles and directions, giving covalently bound molecules distinct, three-dimensional forms. For biological molecules such as proteins, which are held together primarily by covalent bonds, the three-dimensional form imparted by these bonds is critical to their functions.

Carbon, with four unpaired outer electrons, typically forms four covalent bonds to complete its outermost energy level. An example is methane, CH_4 **(Figure 2.8a, b),** the main component of natural gas. The four covalent bonds formed by the carbon atom are fixed at an angle of 109.5° from each other, forming a tetrahedron. The tetrahedral arrangement of the bonds allows carbon "building blocks" **(Figure 2.8c)** to link to each other in both branched and unbranched chains and rings **(Figure 2.8d).** Such structures form the backbones of an almost unlimited variety of molecules. Carbon can also form double bonds, in which atoms share two pairs of electrons, and triple bonds, in which atoms share three pairs of electrons.

Oxygen, hydrogen, nitrogen, and sulfur also share electrons readily to form covalent linkages, and they commonly combine with carbon in biological molecules. In these linkages, oxygen typically forms two covalent bonds; hydrogen, one; nitrogen, three; and sulfur, two.

Unequal Electron Sharing Results in Polarity

Electronegativity is the measure of an atom's attraction for the electrons it shares in a chemical bond with another atom. The more electronegative an atom is, the more strongly it attracts shared electrons. Among atoms, electronegativity increases as the number of protons in the nucleus increases and as the distance of electrons from the nucleus increases.

Although all covalent bonds involve the sharing of valence electrons, they differ widely in the degree of sharing. Depending on the difference in electronegativity between the bonded atoms, the covalent bonds are classified as **nonpolar covalent bonds** or **polar covalent bonds.** In a nonpolar covalent bond, electrons are shared equally, whereas in a polar covalent bond, they are shared unequally. When electron sharing is unequal, as in polar covalent bonds, the atom that attracts the electrons more strongly carries a partial negative charge and the atom deprived of electrons carries a partial positive charge. The atoms carrying partial charges may give the whole molecule partially positive and negative ends; in other words, the molecule is *polar,* hence the name given to the bond.

Nonpolar covalent bonds are characteristic of molecules that contain atoms of one kind, such as hydrogen (H_2) and oxygen (O_2), although there are some exceptions. Polar covalent bonds are characteristic of molecules that contain atoms of different types.

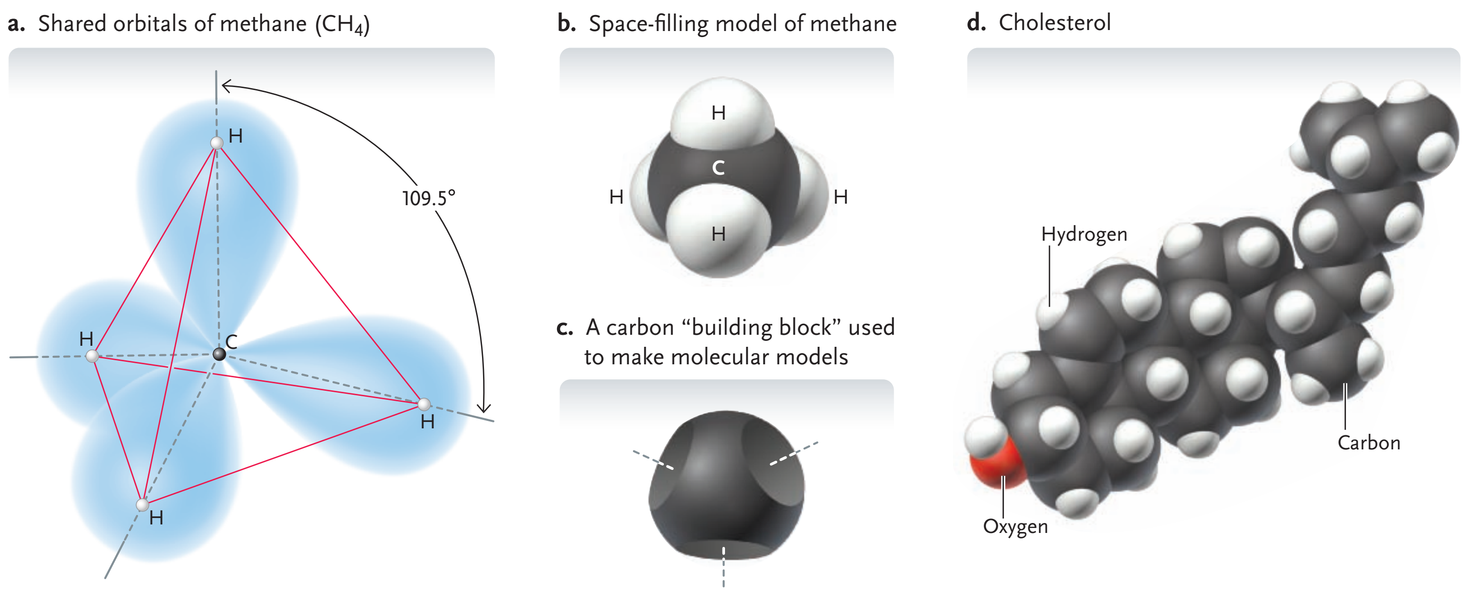


Figure 2.8
Covalent bonds shared by carbon. **(a)** The four covalent bonds of carbon in methane (CH_4) are shown as shared orbitals. The bonds extend outward from the carbon nucleus at angles of 109.5° from each other (dashed lines). The red lines connecting the hydrogen nuclei form a regular tetrahedron with four faces. **(b)** Space-filling model of methane, in which the diameter of the sphere representing an atom shows the approximate limit of its electron orbitals. **(c)** A tetrahedral carbon "building block." One of the four faces of the block is not visible. **(d)** Carbon atoms assembled into rings and chains forming a complex molecule.

For example, in water, an oxygen atom forms polar covalent bonds with two hydrogen atoms. Because the oxygen nucleus with its eight protons attracts electrons much more strongly than the hydrogen nuclei do, the bonds are strongly polar **(Figure 2.9).** In addition, the water molecule is asymmetric, with the oxygen atom located on one side and the hydrogen atoms on the other. This arrangement gives the entire molecule an unequal charge distribution, with the hydrogen end partially positive and the oxygen end partially negative, and makes water molecules strongly polar. In fact, water is the primary biological example of a polar molecule. The polar nature of water underlies its ability to adhere to ions and weaken their attractions.

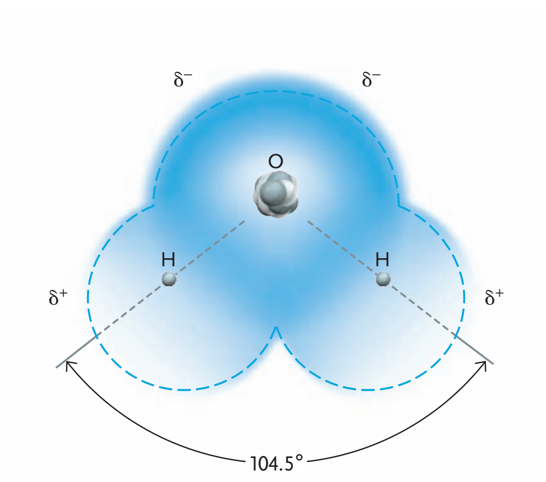


Figure 2.9
Polarity in the water molecule, created by unequal electron sharing between the two hydrogen atoms and the oxygen atom and the asymmetric shape of the molecule. The unequal electron sharing gives the hydrogen end of the molecule a partial positive charge, δ^+ ("delta plus"), and the oxygen end of the molecule a partial negative charge, δ^- ("delta minus"). Regions of deepest color indicate the most frequent locations of the shared electrons. The orbitals occupied by the electrons are more complex than the spherical forms shown here.

Oxygen, nitrogen, and sulfur, which all share electrons unequally with hydrogen, are located asymmetrically in many biological molecules. Therefore, the presence of —OH, —NH, or —SH groups tends to make regions in biological molecules containing them polar.

Although carbon and hydrogen share electrons somewhat unequally, these atoms tend to be arranged symmetrically in biological molecules. Thus, regions that contain only carbon–hydrogen chains are typically nonpolar. For example, the C—H bonds in methane are located symmetrically around the carbon atom (see Figure 2.8), so their partial charges cancel each other and the molecule as a whole is nonpolar.

Polar Molecules Tend to Associate with Each Other and Exclude Nonpolar Molecules

Polar molecules attract and align themselves with other polar molecules and with charged ions and molecules. These **polar associations** create environments that tend to exclude nonpolar molecules. When present in quantity, the excluded nonpolar molecules tend to clump together in arrangements called **nonpolar associations;** these nonpolar associations reduce the surface area exposed to the surrounding polar environment. Polar molecules that associate readily with water are identified as **hydrophilic** (*hydro* = water; *philic* = preferring). Nonpolar substances that are excluded by water and other polar molecules are identified as **hydrophobic** (*phobic* = avoiding).

Polar and nonpolar associations can be demonstrated with an apparatus no more complex than a bottle containing water and vegetable oil. If the bottle has been placed at rest for some time, the nonpolar oil and polar water form separate layers, with the oil on top. If you shake the bottle, the oil becomes suspended as spherical droplets in the water; the harder you shake, the smaller the oil droplets become (the spherical form of the oil droplets exposes the least surface area per unit volume to the watery polar surroundings). If you place the bottle at rest, the oil and water quickly separate again into distinct polar and nonpolar layers.

Hydrogen Bonds Also Involve Unequal Electron Sharing

When hydrogen atoms are made partially positive by sharing electrons unequally with oxygen, nitrogen, or sulfur, they may be attracted to nearby oxygen, nitrogen, or sulfur atoms made partially negative by unequal electron sharing in a different covalent bond **(Figure 2.10a).** This attractive force is the hydrogen bond, illustrated by a dotted line in structural diagrams of molecules. Hydrogen bonds may be intramolecular (between atoms in the same molecule) or intermolecular (between atoms in different molecules).

Individual hydrogen bonds are weak compared with ionic and covalent bonds. However, large biological molecules may offer many opportunities for hydrogen bonding, both within and between molecules. When numerous, hydrogen bonds are collectively strong and lend stability to the three-dimensional structure of molecules such as proteins **(Figure 2.10b).** Hydrogen bonds between water molecules are responsible for many of the properties that make water uniquely important to life (see Section 2.4 for a more detailed discussion).

The weak attractive force of hydrogen bonds makes them much easier to break than covalent and ionic bonds, particularly when elevated temperature increases the movements of molecules. Hydrogen bonds begin to break extensively as temperatures rise above 45°C and become practically nonexistent at 100°C. The disruption of hydrogen bonds by heat—for instance, the bonds in proteins—is one of the primary reasons most organisms cannot survive temperatures much greater than 45°C. Thermophilic (temperature-loving) organisms, which live at temperatures higher than 45°C, some at 120°C or more, have different molecules from those of organisms that live at lower temperatures. For example, proteins in thermophiles are stabilized at high temperatures by van der Waals forces and other noncovalent interactions.

Van der Waals Forces Are Weak Attractions over Very Short Distances

Van der Waals forces are even weaker than hydrogen bonds. These forces develop between nonpolar molecules or regions of molecules when, through their constant motion, electrons accumulate by chance in one part of a molecule or another. This process leads to zones of positive and negative charge, making the molecule polar. If they are oriented in the right way, the polar parts of the molecules are attracted electrically to one another and cause the molecules to stick together briefly. Although an individual bond formed with van der Waals forces is weak and transient, the formation of many bonds of this type can stabilize the shape of a large molecule, such as a protein.

A striking example of the collective power of van der Waals forces concerns the ability of geckos, a group of tropical lizard species, to cling to and walk up vertical smooth surfaces **(Figure 2.11).** The toes of the gecko are covered with millions of hairs, called *setae* (pronounced "see-tea"), that are about 100 micrometers (µm; 0.004 inch) long. At the tip of each hair are hundreds of thousands of pads, each about 200 nanometers (nm;

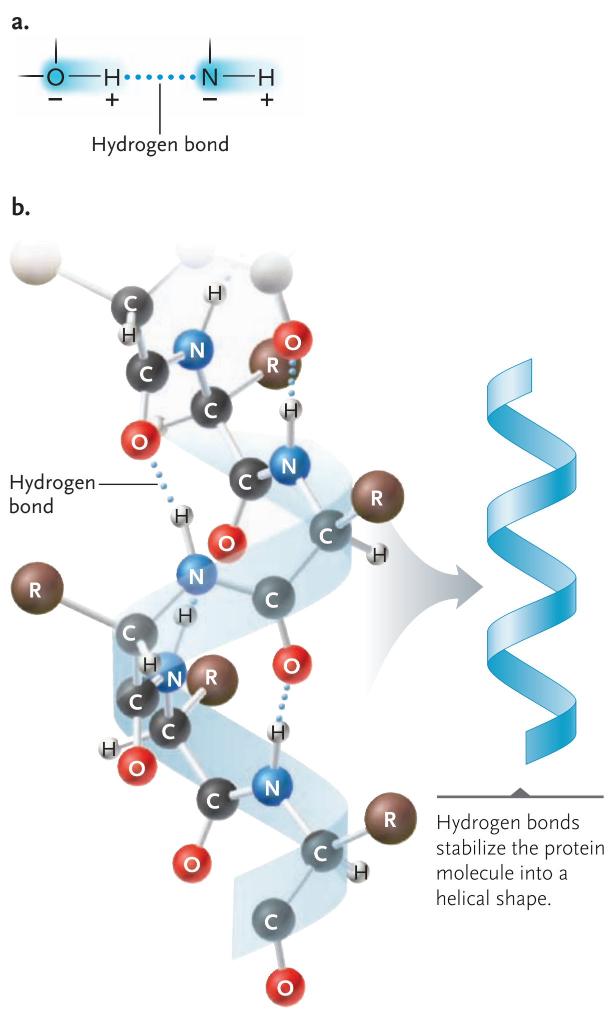
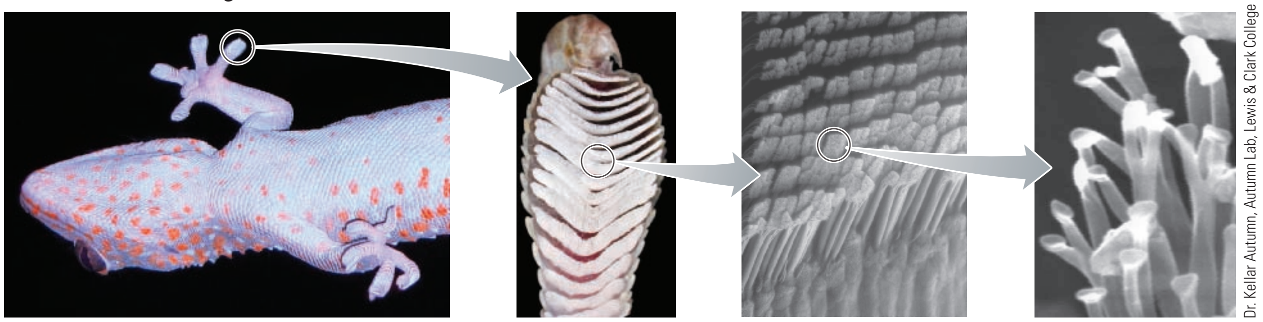


Figure 2.10
Hydrogen bonds. **(a)** A hydrogen bond (dotted line) between the hydrogen of an —OH group and a nearby nitrogen atom, which also shares electrons unequally with another hydrogen. Regions of deepest blue indicate the most likely locations of electrons. **(b)** Multiple hydrogen bonds stabilize the backbone chain of a protein molecule into a spiral called the alpha helix. The spheres labeled *R* represent chemical groups of different kinds.

Figure 2.11 An example of van der Waals forces in biology. **(a)** A Tokay gecko *(Gekko gecko)* climbing while inverted on a glass plate. **(b)** Gecko toe. **(c)** Setae on a toe. **(d)** Pads (spatulae) on an individual seta.



0.000008 inch) wide—smaller than the wavelength of visible light. Each pad forms a weak interaction—using van der Waals forces—with molecules on the surface. Magnified by the huge number of pads involved, the attractive forces are 1000 times greater than necessary for the gecko to hang on a vertical wall. To climb a wall, the animal rolls the hairs onto the surface and then peels them off like a piece of tape. Understanding the gecko's remarkable ability to climb has led to the development of gecko tape, a superadhesive prototype tape capable of holding a 3-kg weight with a 1-cm^2 (centimeter squared) piece.

Bonds Form and Break in Chemical Reactions

Chemical reactions occur when atoms or molecules interact to form new chemical bonds or break old ones. As a result of bond formation or breakage, atoms are added to or removed from molecules, or the linkages of atoms in molecules are rearranged. When any of these alterations occur, molecules change from one type to another, usually with different chemical and physical properties. In biological systems, chemical reactions are accelerated by molecules called *enzymes* (which are discussed in more detail in Chapter 4).

The atoms or molecules entering a chemical reaction are called the **reactants,** and those leaving a reaction are the **products.** A chemical reaction is written with an arrow showing the direction of the reaction; reactants are placed to the left of the arrow, and products are placed to the right. Both reactants and products are usually written in chemical shorthand as formulas.

For example, the overall reaction of photosynthesis, in which carbon dioxide and water are combined to produce sugars and oxygen (see Chapter 9), is written as follows:

$$\underset{\text{carbon dioxide}}{6\ CO_2} + \underset{\text{water}}{6\ H_2O} + \text{light} \rightarrow \underset{\text{a sugar}}{C_6H_{12}O_6} + \underset{\text{molecular oxygen}}{6\ O_2}$$

The number in front of each formula indicates the number of molecules of that type among the reactants and products (the number 1 is not written). Notice that there are as many atoms of each element to the left of the arrow as there are to the right, even though the products are different from the reactants. This balance reflects the fact that in such reactions, atoms may be rearranged but not created or destroyed. Chemical reactions written in balanced form are known as **chemical equations.**

With the information about chemical bonds and reactions provided thus far, you are ready to examine the effects of chemical structure and bonding, particularly hydrogen bonding, in the production of the unusual properties of water, the most important substance to life on Earth.

Study Break

1. Explain how an ionic bond forms.
2. Explain how a covalent bond forms.
3. What is electronegativity, and how does it relate to nonpolar covalent bonds and polar covalent bonds?
4. What is a chemical reaction?

2.4 Hydrogen Bonds and the Properties of Water

All living organisms contain water, and many kinds of organisms live directly in water. Even those that live in dry environments contain water in all their structures—different organisms range from 50% to more than 95% water by weight. The water inside organisms is crucial for life: it is required for many important biochemical reactions and plays major roles in maintaining the shape and organization of cells and tissues. The properties of water molecules that make them so important to life depend to a great extent on their polar structure and their ability to link to each other by hydrogen bonds.

A Lattice of Hydrogen Bonds Gives Water Unusual Properties

Hydrogen bonds form readily between water molecules in both liquid water and ice. In liquid water, each water molecule establishes an average of 3.4 hydrogen bonds with its neighbors, forming an arrangement known as the **water lattice (Figure 2.12a).** In liquid wa-

a. Hydrogen-bond lattice of liquid water

b. Hydrogen-bond lattice of ice

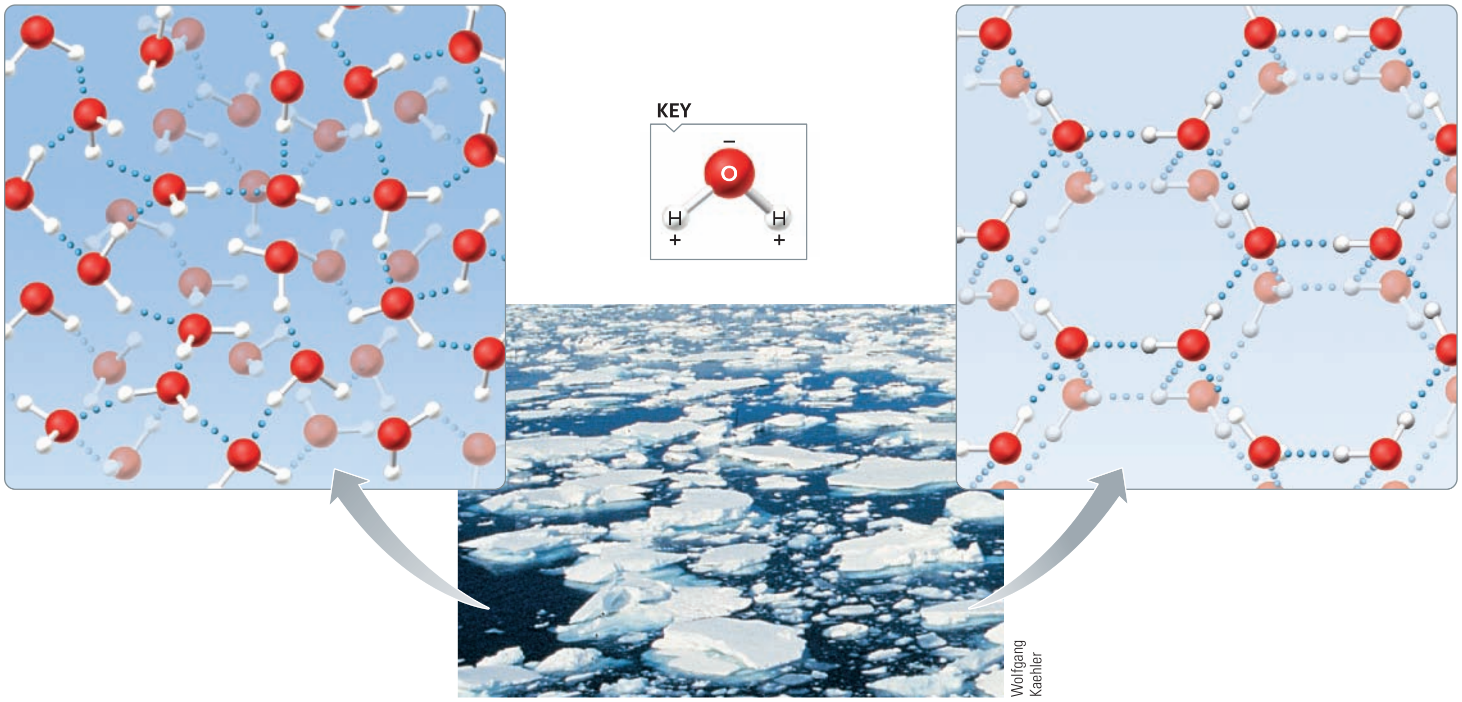


Figure 2.12 Hydrogen bonds and water. **(a)** In liquid water, hydrogen bonds between molecules (dotted lines) form and break rapidly, allowing the molecules to slip past each other easily. **(b)** In ice, water molecules are fixed into a rigid lattice.

ter, the hydrogen bonds that hold the lattice together constantly break and reform, allowing the water molecules to break loose from the lattice, slip past one another, and reform the lattice in new positions.

In ice, the water lattice is a rigid, crystalline structure in which each water molecule forms four hydrogen bonds with neighboring molecules **(Figure 2.12b).** The rigid **ice lattice** spaces the water molecules farther apart than the water lattice. Because of this greater spacing, water has the unusual property of being about 10% less dense when solid than when liquid. (Almost all other substances are denser in solid form than in liquid form.) Hence, ice cubes are a little larger than the water volume poured into the ice tray, and water filling a closed glass vessel will break the vessel when the water freezes. At atmospheric pressure, water reaches its greatest density at a temperature of 4°C, while it is still a liquid.

Because it is less dense than liquid water, ice forms at the surface of a body of water and remains floating at the surface. The ice creates an insulating layer that helps keep the water below from freezing. If ice were denser than liquid water, it would sink to the bottom as it freezes, continually exposing liquid water at the surface to freezing. Under those conditions, most bodies of water would freeze entirely solid, making life difficult or impossible for aquatic plants and animals.

The Hydrogen-Bond Lattice of Water Contributes to Polar and Nonpolar Environments in and around Cells

The hydrogen-bond lattice and the polarity of water molecules give water other properties that make it unique and ideal as a life-sustaining medium. In liquid water, the lattice resists invasion by other molecules unless the invading molecule also contains polar or charged regions that can form competing attractions with water molecules. If present, the competing attractions open the water lattice, creating a cavity into which the polar or charged molecule can move. By contrast, nonpolar molecules are unable to disturb the water lattice. The lattice thus excludes nonpolar substances, forcing them to form the nonpolar associations that expose the least surface area to the surrounding water—such as the spherical droplets of oil that form when oil and water are shaken.

The distinct polar and nonpolar environments created by water are critical to the organization of cells. For example, biological membranes, which form boundaries around and inside cells, consist of lipid molecules with dual polarity: one end of each molecule is polar, and the other end is nonpolar. (Lipids are described in more detail in Chapter 3.) The membranes are surrounded on both sides by strongly polar water molecules. Exclusion by the water molecules forces the lipid molecules to associate into a double layer, a **bilayer,** in which only the polar ends of the surface molecules are exposed to the water **(Figure 2.13).** The nonpolar ends of the molecules associate in the interior of the bilayer, where they are not exposed to the water. Exclusion of their nonpolar regions by water is all that holds membranes together.

The membrane at the surface of cells prevents the watery solution inside the cell from mixing directly with the watery solution outside the cell. By doing so, the surface membrane, kept intact by nonpolar exclusion by water, maintains the internal environment and organization necessary for cellular life.

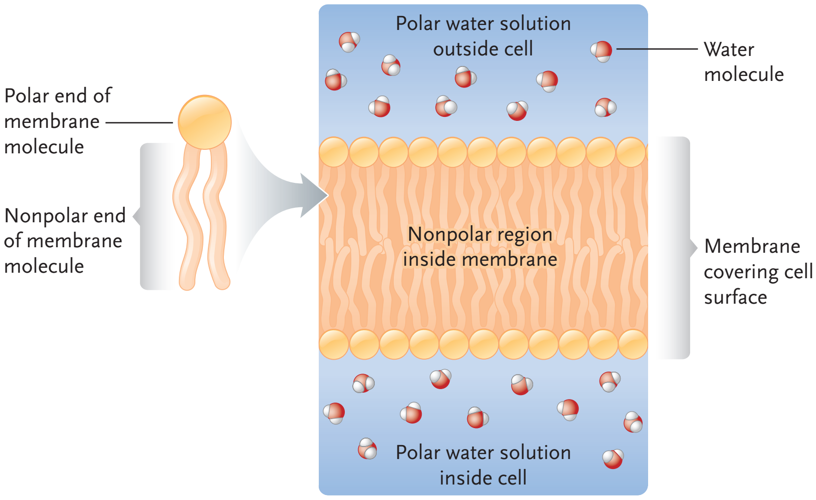


Figure 2.13
Double layer of lipid molecules that forms a membrane covering the cell surface.
Exclusion by polar water molecules forces the nonpolar ends of the surface molecules to associate into the thin, double layer—the bilayer—that forms the membrane.

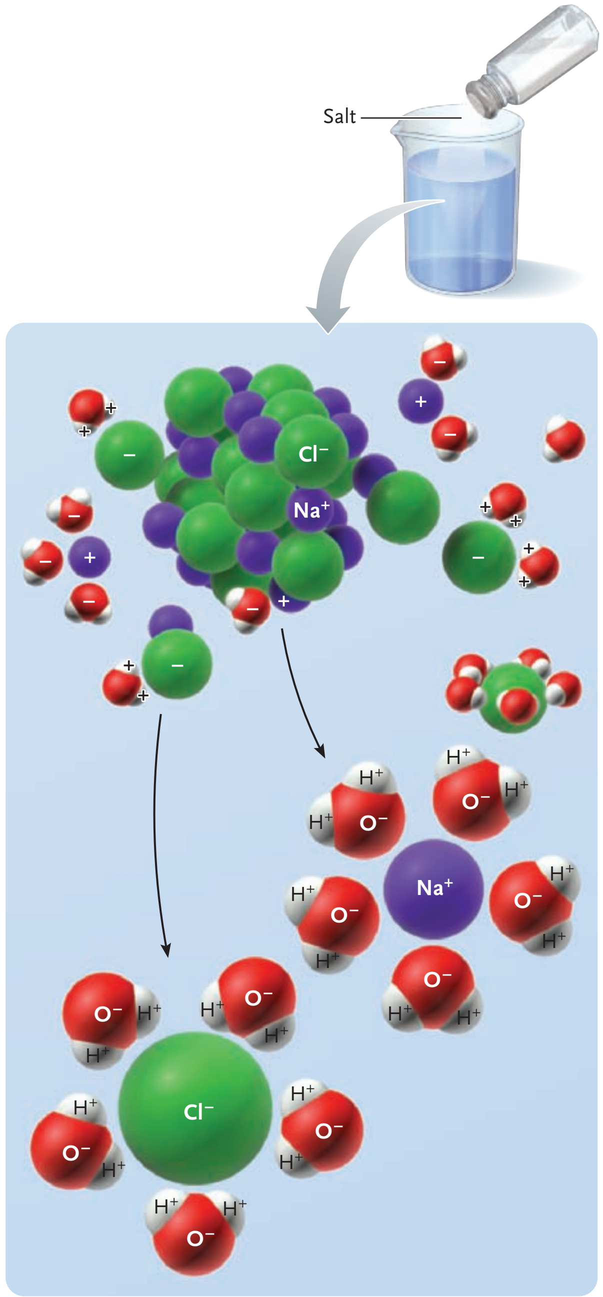


Figure 2.14
Water molecules forming a hydration layer around Na^+ and Cl^- ions, which promotes their separation and entry into solution.

The Small Size and Polarity of Its Molecules Makes Water a Good Solvent

Because water molecules are small and strongly polar, they readily penetrate or coat the surfaces of other polar and charged molecules and ions. The surface coat, called a **hydration layer,** reduces the attraction between the molecules or ions and promotes their separation and entry into a **solution,** where they are suspended individually, surrounded by water molecules. Once in solution, the hydration layer prevents the polar molecules or ions from reassociating. In such a solution, water is called the **solvent,** and the molecules of a substance dissolved in water are called the **solute.**

For example, when a teaspoon of table salt is added to water, water molecules quickly form hydration layers around the Na^+ and Cl^- ions in the salt crystals, reducing the attraction between the ions so much that they separate from the crystal and enter the surrounding water lattice as individual ions **(Figure 2.14).** If the water evaporates, the hydration layer is eliminated, exposing the strong positive and negative charges of the ions. The opposite charges attract and reestablish the ionic bonds that hold the ions in salt crystals. As the last of the water evaporates, all of the Na^+ and Cl^- ions relink into the solid, crystalline form.

In the cell, chemical reactions depend on solutes dissolved in aqueous solutions. To understand these reactions, you need to know the number of atoms and molecules involved. **Concentration** is the number of molecules or ions of a substance in a unit volume of space, such as a milliliter (mL) or liter (L). The number of molecules or ions in a unit volume cannot be counted directly but can be calculated indirectly by using the mass number of atoms as the starting point. The same method is used to prepare a solution with a known number of molecules per unit volume.

The mass number of an atom is equivalent to the number of protons and neutrons in its nucleus. From the mass number, and the fact that neutrons and protons are approximately the same weight (that is, 1.66×10^{-24} g), you can calculate the weight of an atom of any substance. For an atom of the most common form of carbon, with 6 protons and 6 neutrons in its nucleus, the total weight is:

$$12 \times (1.66 \times 10^{-24}\text{ g}) = 1.992 \times 10^{-23}\text{ g}$$

For an oxygen atom, with 8 protons and 8 neutrons in its nucleus, the total weight is:

$$16 \times (1.66 \times 10^{-24}\text{ g}) = 2.656 \times 10^{-23}\text{ g}$$

Dividing the total weight of a sample of an element by the weight of a single atom gives the number of atoms in the sample. Suppose you have a carbon sample that weighs 12 g—a weight in grams equal to the atom's mass number. (A weight in grams equal to the mass number is known as an **atomic weight** of an element.) Dividing 12 g by the weight of one carbon atom gives:

$$\frac{12}{(1.992 \times 10^{-23}\,\mathrm{g})} = 6.022 \times 10^{23} \text{ atoms}$$

If you divide the atomic weight of oxygen (16 g) by the weight of one oxygen atom, you get the same result:

$$\frac{16}{(2.656 \times 10^{-23}\,\mathrm{g})} = 6.022 \times 10^{23} \text{ atoms}$$

In fact, dividing the atomic weight of any element by the weight of an atom of that element always produces the same number: 6.022×10^{23}. This number is called **Avogadro's number** after Amedeo Avogadro, the nineteenth-century Italian chemist who first discovered the relationship.

The same relationship holds for molecules. The **molecular weight** of any molecule is the weight in grams equal to the total mass number of its atoms. For NaCl, the total mass number is 23 + 35 = 58 (a sodium atom has 11 protons and 12 neutrons, and a chlorine atom has 17 protons and 18 neutrons). The weight of an NaCl molecule is therefore:

$$58 \times (1.66 \times 10^{-24}\,\mathrm{g}) = 9.628 \times 10^{-23}\,\mathrm{g}$$

Dividing a molecular weight of NaCl (58 g) by the weight of a single NaCl molecule gives:

$$\frac{58}{(9.628 \times 10^{-23}\,\mathrm{g})} = 6.022 \times 10^{23} \text{ molecules}$$

When concentrations are described, the atomic weight of an element or the molecular weight of a compound—the amount that contains 6.022×10^{23} atoms or molecules—is known as a **mole** (abbreviated **mol**). The number of moles of a substance dissolved in 1 L of solution is known as the **molarity** (abbreviated ***M***) of the solution. This relationship is highly useful in chemistry and biology because we know that two solutions having the same volume and molarity but composed of different substances will contain the same number of molecules of the substances.

The Hydrogen-Bond Lattice Gives Water Other Life-Sustaining Properties as Well

The hydrogen-bond lattice gives water other unique properties that make it a medium suitable for the molecules and reactions of life. Compared with substances that have a similar molecular structure, such as H_2S (hydrogen sulfide):

- Water has an unusual ability to resist changes in temperature by absorbing or releasing heat, plus an unusually high boiling point.
- Water has an unusually high internal cohesion and surface tension.

The Boiling Point and Temperature-Stabilizing Effects of Water. The hydrogen-bond lattice of liquid water retards the escape of individual water molecules as the water is heated. As a result, relatively high temperatures and the addition of considerable heat are required to break enough hydrogen bonds to make water boil. The high boiling point maintains water as a liquid over the wide temperature range of 0° to 100°C. Similar molecules that do not form an extended hydrogen-bond lattice, such as H_2S, have much lower boiling points and are gases rather than liquids at room temperature. The properties of these related substances indicate that without its hydrogen-bond lattice, water would boil at −81°C. If this were the case, most of the water on Earth would be in gaseous form and life as described in this book could not exist.

As a result of water's stabilizing hydrogen-bond lattice, it also has a relatively high **specific heat**—that is, the amount of heat required to increase the temperature of a given quantity of water. As heat flows into water, much of it is absorbed in the breakage of hydrogen bonds. As a result, the temperature of water, reflected in the average motion of its molecules, increases relatively slowly as heat is added. For example, a given amount of heat increases the temperature of water by only half as much as that of an equal quantity of ethyl alcohol. High specific heat allows water to absorb or release relatively large quantities of heat without undergoing extreme changes in temperature; this gives it a moderating and stabilizing effect on both living organisms and their environments.

The specific heat of water is measured in **calories.** This unit, used both in the sciences and in dieting, is the amount of heat required to raise 1 g of water by 1°C (technically, from 14.5 to 15.5°C at one atmosphere of pressure). This amount of heat is known as a "small" calorie and is written with a small *c*. The unit most familiar to dieters, equal to 1000 small calories, is written with a capital *C* as a **Calorie;** the same 1000-calorie unit is known scientifically as a **kilocalorie (kcal).** A 300-Calorie candy bar therefore really contains 300,000 calories.

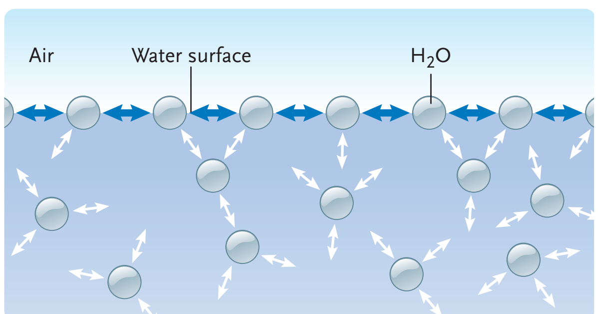
A large amount of heat, 586 calories per gram, must be added to give water molecules enough energy of motion to break loose from liquid water and form a gas. This required heat, known as the **heat of vaporization,** allows humans and many other organisms to cool off when hot. In humans, water is released onto the surface of the skin by more than 2.5 million sweat glands; the heat absorbed by this water as it evaporates cools the skin and the underlying blood vessels. The heat loss helps keep body temperature from increasing when environmental temperatures are high. Plants use a similar cooling mechanism as water evaporates from their leaves.

Cohesion and Surface Tension. The high resistance of water molecules to separation, provided by the hydrogen-bond lattice, is known as internal **cohesion.** For example, in land plants, cohesion holds water molecules in unbroken columns in microscopic conducting tubes that extend from the roots to the highest leaves. As water evaporates from the leaves, water molecules in the columns, held together by cohesion, move upward through the tubes to replace the lost water. This movement raises water from roots to the tops of the tallest trees (see discussion in Chapter 32). Maintenance of the long columns of water in the tubes is aided by **adhesion,** in which molecules "stick" to the walls of the tubes by forming hydrogen bonds with charged and polar groups in molecules that form the walls of the tubes.

Water molecules at surfaces facing air can form hydrogen bonds with water molecules beside and below them but not on the sides that face the air. This unbalanced bonding produces a force that places the surface water molecules under tension, making them more resistant to separation than the underlying water molecules **(Figure 2.15a).** The force, called **surface tension,** is strong enough to allow small insects such as water striders to walk on water **(Figure 2.15b).** Similarly, the surface tension of water will support a sewing needle placed carefully on the surface, even though the needle is about 10 times denser than the water. Surface tension also causes water to form water droplets; the surface tension pulls the water in around itself to produce the smallest possible area, which is a spherical bead or droplet.

Water has still other properties that contribute to its ability to sustain life, the most important being that its molecules separate into ions. These ions help maintain an environment inside living organisms that promotes the chemical reactions of life.

a.



b.

Figure 2.15
Surface tension in water. **(a)** The unbalanced hydrogen bonding that places water molecules under lateral tension where a water surface faces the air. **(b)** A water strider (*Gerris* species) supported by the surface tension of water.

Study Break

1. How do hydrogen bonds between water molecules contribute to the properties of water?
2. Distinguish between a solute, a solvent, and a solution.

2.5 Water Ionization and Acids, Bases, and Buffers

The most critical property of water that is unrelated to its hydrogen-bond lattice is its ability to separate, or **dissociate**, to produce positively charged *hydrogen ions* (H^+, or protons) and *hydroxide ions* (OH^-):

$$H_2O \rightleftharpoons H^+ + OH^-$$

(The double arrow means that the reaction is **reversible**—that is, depending on conditions, it may go from left to right or from right to left.) The proportion of water molecules that dissociates to release protons and hydroxide ions is small. However, because of the dissociation, water always contains some H^+ and OH^- ions.

Substances Act as Acids or Bases by Altering the Concentrations of H^+ and OH^- Ions in Water

In pure water, the concentrations of H^+ and OH^- ions are equal. However, adding other substances may alter the relative concentrations of H^+ and OH^-, making them unequal. Some substances, called **acids,** are proton donors that release H^+ (and anions) when they are dissolved in water, effectively increasing the H^+ concentration. For example, hydrochloric acid (HCl) dissociates into H^+ and Cl^- when dissolved in water:

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

Other substances, called **bases,** are proton acceptors that reduce the H^+ concentration of a solution. Most bases dissociate in water into a hydroxide ion (OH^-) and a cation. The hydroxide ion can act as a base by accepting a proton (H^+) to produce water. For example, sodium hydroxide (NaOH) separates into Na^+ and OH^- ions when dissolved in water:

$$NaOH \rightarrow Na^+ + OH^-$$

The excess OH^- combines with H^+ to produce water:

$$OH^- + H^+ \rightarrow H_2O$$

thereby reducing the H^+ concentration.

Other bases do not dissociate to produce hydroxide ions directly. For example, ammonia (NH_3), a poisonous gas, acts as a base when dissolved in water, directly accepting a proton from water to produce an ammonium ion and releasing a hydroxide ion:

$$NH_3 + H_2O \rightarrow NH_4{}^+ + OH^-$$

The concentration of H^+ ions in a water solution, as compared with the concentration of OH^- ions, determines the **acidity** of the solution. Scientists measure acidity using a numerical scale from 0 to 14, called the **pH scale.** Because the number of H^+ ions in solution increases exponentially as the acidity increases, the scale is based on logarithms of this number to make the values manageable:

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

In this formula, the brackets indicate concentration in moles per liter. The negative of the logarithm is used to give a positive number for the pH value. For example, in a water solution that is *neutral*—neither acidic nor basic—the concentration of *both* H^+ and OH^- ions is 1×10^{-7} *M* (0.0000001 *M*). The $\log_{10}$ of 1×10^{-7} is -7. The negative of the logarithm -7 is 7. Thus, a neutral water solution with an H^+ concentration of 1×10^{-7} *M* has a pH of 7. *Acidic* solutions have pH values less than 7, with pH 0 being the value for the highly acidic 1 *M* hydrochloric acid (HCl); *basic* solutions have pH values greater than 7, with pH 14 being the value for the highly basic 1 *M* sodium hydroxide (NaOH) (basic solutions are also called *alkaline* solutions). Each whole number on the pH scale represents a value 10 times greater or less than the next number. Thus, a solution with a pH of 4 is 10 times more acidic than one with a pH of 5, and a solution with a pH of 6 is 100 times more acidic than a solution with a pH of 8. (The pH of many familiar solutions is shown in **Figure 2.16.**)

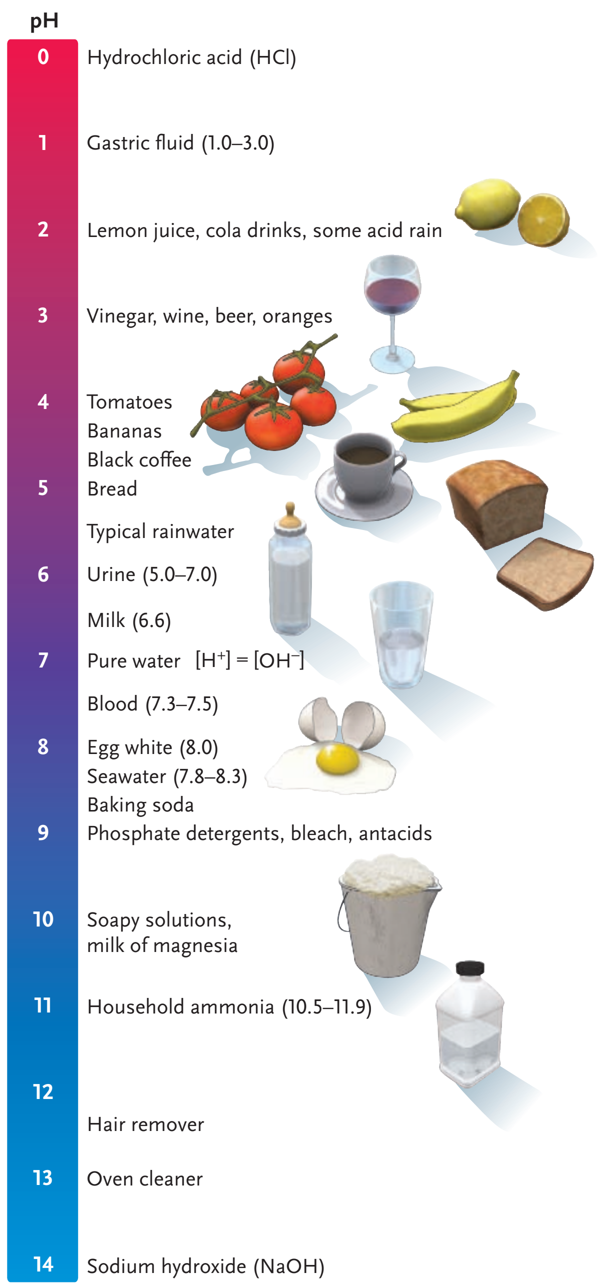


Figure 2.16
The pH scale, showing the pH of substances commonly encountered in the environment.

Acidity is important to cells because even small changes, on the order of 0.1 or even 0.01 pH unit, can drastically affect biological reactions. In large part, this effect reflects changes in the structure of proteins that occur when the water solution surrounding the proteins has too few or too many hydrogen ions. Consequently, all living organisms have elaborate systems that control their internal acidity by regulating H^+ concentration near the neutral value of pH 7.

Acidity is also important to the environment in which we live. Where the air is unpolluted, rainwater is only slightly acidic. However, in regions where certain pollutants are released into the air in large quantities by industry and automobile exhaust, the polluting chemicals combine with atmospheric water to produce "acid rain" with a pH as low as 3, about the same pH as that of vinegar. Acid rain can sicken and kill wildlife such as fishes and birds, as well as plants and trees (**Figure 2.17;** see also discussion in Chapter 53). Humans are also affected; acid rain and acidified water vapor in the air can contribute to human respiratory diseases such as bronchitis and asthma.

Buffers Help Keep pH under Control

Living organisms control the internal pH of their cells with **buffers**, substances that compensate for pH changes by absorbing or releasing H^+. When H^+ ions are released in excess by biological reactions, buffers combine

Frederica Georgia/Photo Researchers, Inc.

Figure 2.17
Forest affected by acid rain and other forms of air pollution in the Great Smoky Mountains National Park. The trees are susceptible to drought, disease, and insect pests.

with them and remove them from the solution; if the concentration of H^+ decreases greatly, buffers release additional H^+ to restore the balance. Most buffers are weak acids or bases, or combinations of these substances, that dissociate reversibly in water solutions to release or absorb H^+ or OH^-. (Weak acids, such as acetic acid, or weak bases, such as ammonia, are substances that release relatively few H^+ or OH^- ions in a water solution. Strong acids or bases are substances that dissociate extensively in a water solution. HCl is a strong acid; NaOH is a strong base.)

The buffering mechanism that maintains blood pH near neutral values is a primary example. In humans and many other animals, blood pH is buffered by a chemical system based on carbonic acid (H_2CO_3), a weak acid. In water solutions, carbonic acid dissociates readily into bicarbonate ions (HCO_3^-) and H^+:

$$H_2CO_3 \rightleftharpoons HCO_3^- + H^+$$

The reaction is reversible. If H^+ is present in excess, the reaction is pushed to the left—the excess H^+ ions combine with bicarbonate ions to form H_2CO_3. If the H^+ concentration declines below normal levels, the reaction is pushed to the right—H_2CO_3 dissociates into HCO_3^- and H^+, restoring the H^+ concentration. The back-and-forth adjustments of the buffer system help keep human blood close to its normal pH of 7.4.

The effects of hyperventilation highlight the importance of the system that buffers blood pH. Hyperventilation, caused by breathing too fast, drastically reduces the CO_2 concentration in blood **(Figure 2.18).** Carbon dioxide is the primary source of carbonic acid in the bloodstream ($CO_2 + H_2O \rightarrow H_2CO_3$); removing too much CO_2 causes the amount of carbonic acid in the blood to decrease. If the amount of blood CO_2 drops so low that the carbonic acid buffer is no longer able to maintain pH at normal levels, a series of internal reac-

Unanswered Questions

Can arsenic be removed from the soil by bioremediation?

In the Why It Matters section, we learned that bioremediation of selenium in wastewater is possible using plants. Research is showing that bioremediation can also be used to remove other toxic chemicals in the environment, including perchlorate and arsenic. For example, arsenic-contaminated soils and sediments are the major sources of arsenic contamination in surface water and groundwater, which leads to contamination of foods. In some parts of the world, the drinking water is contaminated. Arsenic poses serious health risks to humans and other animals; for example, some cancers have been correlated with high levels of arsenic. Arsenic contamination is a worldwide concern, with arsenic levels in the environment in some parts of the world being tens of thousands of times higher than the maximum contaminant level set in the United States.

One research group at LaTrobe University, Melbourne, Australia, led by Joanne Santini, is exploring whether bacteria can be used for arsenic bioremediation in contaminated wastewater on mining sites and from groundwater in Bangladesh and West Bengal, India. Their approach has been to study 13 rare bacteria isolated from gold mines, a typical place to find arsenic. Arsenic is present in water in two toxic forms; one of these forms is easy and safe to get rid of, but the other is not. Santini's group has identified a bacterium that can "eat" the difficult-to-get-rid-of form of arsenic and convert it to the easy-to-get-rid-of form. Potentially, this bacterium could be developed for use in bioremediation of arsenic in contaminated locations.

Is food irradiation effective for killing microorganisms?

Radioisotopes are widely used to answer questions in biological research and as tools in medicine. Radioisotopes are also used to irradiate foods with the goal of killing microorganisms capable of causing disease. In most instances, the irradiation of food is done using the radioactive element cobalt-60 as a source of high-energy gamma rays. The energy of the gamma rays is sufficient to dislodge electrons from some food molecules, converting them to ions. But, there is insufficient energy to affect the neutrons in the nuclei of those molecules, so the food is not rendered radioactive by the treatment.

The effectiveness of food irradiation is tested in the laboratory. Researchers perform experiments to determine the dosage needed to kill a population of various pathogens in food. They have shown that irradiation kills many bacteria and parasites and destroys some viruses in food; moreover, they have not seen the development of radiation resistance in the microbial strains and species tested. However, some viruses and spore-forming bacteria are not destroyed by irradiation.

While many organizations such as the World Health Organization (WHO), U.S. Food and Drug Administration (FDA), and Institute of Food Science and Technology have concluded that irradiation of food is safe and can be effective in killing microbial contaminants, questions remain in some quarters, including with some consumers. For example, Does irradiation destroy vitamins? and Are toxic products produced by irradiation?

Researchers have shown that although vitamins in solution can be degraded by irradiation, they are less sensitive to irradiation when present in the complex chemical organization of food. There is some evidence, though, that irradiation sometimes causes chemical changes in food similar to those produced during cooking. Evidence from studies with laboratory animals indicated no adverse health effects when irradiated foods containing these compounds were consumed. However, some concerns remain about the generation of potentially harmful chemical compounds if the food is irradiated in its final packaging. More research needs to be done to determine how serious this concern is to human health.

Peter J. Russell

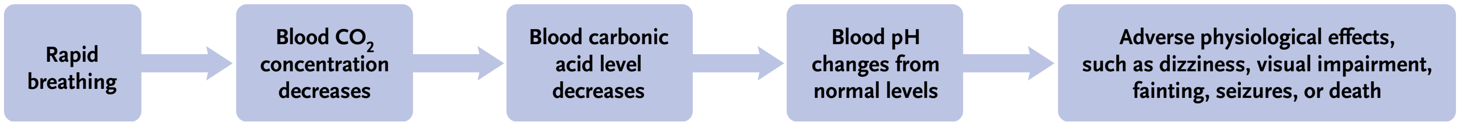


Figure 2.18 Effects of hyperventilation.

tions occurs that can produce dizziness, visual impairment, fainting, seizures, or even death.

This chapter examined the basic structure of atoms and molecules and discussed the unusual properties of water that make it ideal for supporting life. The next chapter looks more closely at the structure and properties of carbon and at the great multitude of molecules based on this element.

Study Break

1. Distinguish between acids and bases. What are their properties?
2. Why are buffers important for living organisms?

Review

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2.1 The Organization of Matter: Elements and Atoms

- Matter is anything that occupies space and has mass. Matter is composed of elements, each consisting of atoms of the same kind.
- Atoms combine chemically in fixed numbers and ratios to form the molecules of living and nonliving matter. Compounds are molecules in which the component atoms are different.

2.2 Atomic Structure

- Atoms consist of an atomic nucleus that contains protons and neutrons surrounded by one or more electrons traveling in orbitals. Each orbital can hold a maximum of two electrons (Figure 2.3).
- All atoms of an element have the same number of protons, but the number of neutrons is variable. The number of protons in an atom is designated by its atomic number; the number of protons plus neutrons is designated by the mass number (Figure 2.4 and Table 2.1).
- Isotopes are atoms of an element with differing numbers of neutrons. The isotopes of an atom differ in physical but not chemical properties (Figure 2.4).
- Electrons surround an atomic nucleus in orbitals occupying energy levels that increase in discrete steps (Figures 2.5 and 2.6).
- The chemical activities of atoms are determined largely by the number of electrons in the outermost energy level. Atoms that have the outermost level filled with electrons are nonreactive, whereas atoms in which that level is not completely filled with electrons are reactive. Atoms tend to lose, gain, or share electrons to fill the outermost energy level.

Video: Isotopes of hydrogen

Animation: Electron arrangements in atoms

Animation: The shell model of electron distribution

Practice: Predicting the number of bonds of elements

2.3 Chemical Bonds

- An ionic bond forms between atoms that gain or lose electrons in the outermost energy level completely, that is, between a positively charged cation and a negatively charged anion (Figure 2.7).
- A covalent bond is established by a pair of electrons shared between two atoms. If the electrons are shared equally, the covalent bond is nonpolar (Figure 2.8).
- If electrons are shared unequally in a covalent bond, the atoms carry partial positive and negative charges and the bond is polar (Figure 2.9).
- Polar molecules tend to associate with other polar molecules and to exclude nonpolar molecules. Polar molecules that associate readily with water are hydrophilic; nonpolar molecules excluded by water are hydrophobic.
- A hydrogen bond is a weak attraction between a hydrogen atom made partially positive by unequal electron sharing and another atom—usually oxygen, nitrogen, or sulfur—made partially negative by unequal electron sharing (Figure 2.10).
- Van der Waals forces, bonds even weaker than hydrogen bonds, can form when natural changes in the electron density of molecules produce regions of positive and negative charge, which cause the molecules to stick together briefly.
- Chemical reactions occur when molecules form or break chemical bonds. The atoms or molecules entering into a chemical reaction are the reactants, and those leaving a reaction are the products.

Animation: How atoms bond

2.4 Hydrogen Bonds and the Properties of Water

- The hydrogen-bond lattice formed by polar water molecules makes it difficult for nonpolar substances to penetrate the lattice. The distinct polar and nonpolar environments created by water are critical to the organization of cells (Figures 2.12 and 2.13).
- The polar properties of water allow it to form a hydration layer over the surfaces of polar and charged biological molecules, particularly proteins. Many chemical reactions depend on the special molecular conditions created by the hydration layer (Figure 2.14).
- The polarity of water allows ions and polar molecules to dissolve readily in water, making it a good solvent.

- The hydrogen-bond lattice gives water unusual properties that are vital to living organisms, including high specific heat, boiling point, cohesion, and surface tension (Figure 2.15).

  Animation: Structure of water

  Animation: Spheres of hydration

2.5 Water Ionization and Acids, Bases, and Buffers

- Acids are substances that increase the H^+ concentration by releasing additional H^+ as they dissolve in water; bases are substances that decrease the H^+ concentration by gathering H^+ or releasing OH^- as they dissolve.
- The relative concentrations of H^+ and OH^- in a water solution determine the acidity of the solution, which is expressed quantitatively as pH on a number scale ranging from 0 to 14. Neutral solutions, in which the concentrations of H^+ and OH^- are equal, have a pH of 7. Solutions with pH less than 7 have H^+ in excess and are acidic; solutions with pH greater than 7 have OH^- in excess and are basic or alkaline (Figure 2.16).
- The pH of living cells is regulated by buffers, which absorb or release H^+ to compensate for changes in H^+ concentration.

  Animation: The pH scale

Questions

Self-Test Questions

1. Which of the following statements about the mass number of an atom is *incorrect?*
   a. It has a unit defined as a dalton.
   b. On Earth, it equals the atomic weight.
   c. Unlike the atomic weight of an atom, it does not change when gravitational forces change.
   d. It equals the number of electrons in an atom.
   e. It is the sum of the protons and neutrons in the atomic nucleus.
2. To make 5 L of a 0.2 *M* aqueous solution of glucose, how many grams of glucose ($C_6H_{12}O_6$) do you need? Atomic masses are carbon 12, hydrogen 1, oxygen 16.
   a. 18.1 b. 180 c. 181 d. 905 e. 9.05
3. The chemical activity of an atom:
   a. depends on the electrons in the outermost energy level.
   b. is increased when the outermost energy level is filled with electrons.
   c. depends on its 1*s* but not its 2*s* or 2*p* orbitals.
   d. is increased when valence electrons completely fill the outer orbitals.
   e. of oxygen prevents it from sharing its electrons with other atoms.
4. When electrons are shared equally, this represents a (an):
   a. polar covalent bond.
   b. nonpolar covalent bond.
   c. ionic bond.
   d. hydrogen bond.
   e. van der Waals force.
5. Which of the following is *not* a property of water?
   a. It has a low boiling point compared with other molecules.
   b. It has a high heat of vaporization.
   c. Its molecules resist separation, a property called cohesion.
   d. It has the property of adhesion, the ability to stick to charged and polar groups in molecules.
   e. It can hydrogen bond to molecules below but not above its surface.
6. Which of the following would *not* represent a hydrophilic body fluid?
   a. blood
   b. sweat
   c. tears
   d. oil
   e. saliva
7. The water lattice:
   a. is formed from hydrophobic bonds.
   b. causes ice to be denser than water.
   c. reduces water's ability as a solvent.
   d. excludes polar substances.
   e. contributes to polar and nonpolar spaces around cells.
8. A hydrogen bond is:
   a. a strong attraction between hydrogen and another atom.
   b. a bond between a hydrogen atom already covalently bound to one atom and made partially negative by unequal electron sharing with another atom.
   c. a bond between a hydrogen atom already covalently bound to one atom and made partially positive by unequal electron sharing with another atom.
   d. weaker than van der Waals forces.
   e. exemplified by the two hydrogens covalently bound to oxygen in the water molecule.
9. If the water in a pond has a pH of 5, the hydroxide concentration would be
   a. 10^{-5} *M*.
   b. 10^{-10} *M*.
   c. 10^{5} *M*.
   d. 10^{9} *M*.
   e. 10^{-9} *M*.
10. Because of a sudden hormonal imbalance, a patient's blood was tested and shown to have a pH of 7.46. What does this pH value mean?
    a. This is more acidic than normal blood.
    b. It represents a weak alkaline fluid.
    c. This is caused by a release of large amounts of hydrogen ions into the system.
    d. The reaction $H_2CO_2 \rightarrow HCO_3^- + H^+$ is pushed to the left.
    e. This is probably caused by excess CO_2 in the blood.

Questions for Discussion

1. Detergents allow particles of oil to mix with water. From the information presented in this chapter, how do you think detergents work?
2. What would living conditions be like on Earth if ice were denser than liquid water?
3. You place a metal pan full of water on the stove and turn on the heat. After a few minutes, the handle is too hot to touch but the water is only warm. How do you explain this observation?
4. You are studying a chemical reaction accelerated by an enzyme. H^+ forms during the reaction, but the enzyme's activity is lost at low pH. What could you include in the reaction mix to keep the enzyme's activity at high levels? Explain how your suggestion might solve the problem.

Experimental Analysis

You know that adding NaOH to HCl results in the formation of common table salt, NaCl. You have a 0.5 *M* HCl solution. What weight of NaOH would you need to add to convert all of the HCl to NaCl? (**Note:** Chemical reactions have the potential to be dangerous. Please do not attempt to perform this reaction.)

Evolution Link

What properties of water made the evolution of life possible?