

# The Cell

**W**e're going to take a fairly long journey through the various levels of biological organization from atoms to ecosystems, as shown in Figure 1.2. Whenever you get ready to go on a trip, you think about what you should bring with you and how to pack your suitcase. Similarly, you can think of the chapters in Part I as the necessities you are going to bring with you as we take our biological journey.

The chapters in Part I will teach you certain principles of biology that will apply to every chapter in the book. Chapters 2 and 3 introduce you to chemistry because all organisms are composed of chemicals, some of them quite unique to living things. In Chapters 4 and 5, we will see how these chemicals are arranged to form the structure of a cell, the basic unit of life. Some organisms are single cells and some are multicellular, but all are made up of cells. Chapters 6, 7, and 8 are about the physiology of cells—how they stay alive.

If these chapters are well understood, you will happily launch forth to study the other parts of the book, secure in the knowledge that you have left nothing behind.

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# 2

## Basic Chemistry

**C**an we understand a bottle-nosed dolphin, without a fundamental knowledge and respect for its chemistry? After all, a dolphin has a certain salinity tolerance, can only stay underwater for so long, and must have a particular diet to keep its complex organ systems functioning. Chemistry also plays a role in the behavior of the dolphin, whether it is playing in the Gulf of Mexico or performing at Sea World. A dolphin cannot jump unless its nervous system is prepared to chemically direct its muscles to contract. In fact, all aspects of a dolphin's biology involve molecular chemistry.

At one time, it was believed that organisms contained a vital force, and this force accounted for their "vitality." Such a hypothesis has never been supported, and instead, today we know that living things are composed of the same elements as inanimate objects. It is true, though, that they differ as to which elements are most common, as we shall see. This chapter reviews inorganic chemistry, which largely pertains to nonliving things. It also explores the composition and chemistry of water, an inorganic substance that is so intimately connected to the life of organisms.

Bottle-nosed dolphin, *Tursiops truncatus*.



### 2.1 CHEMICAL ELEMENTS

- All matter is made up of elements, each of which contains atoms of a particular type. Subatomic particles, particularly the electrons, determine the characteristics of an atom. 22–25

### 2.2 COMPOUNDS AND MOLECULES

- Atoms react with one another to form compounds, whose smallest unit is a molecule. Ionic bonding or covalent bonding can occur between the atoms of a molecule. 26–27

### 2.3 CHEMISTRY OF WATER

- The properties of water make life, as we know it, possible. Hydrogen bonding accounts for why water has a high heat capacity, water is cohesive and adhesive, and frozen water is less dense than liquid water. 28–31

### 2.4 ACIDS AND BASES

- Water, acids, and bases differ by the number of hydrogen ions and hydroxide ions they contain. The pH scale is used to designate the acidity and basicity of a solution. 32–33



## 2.1 Chemical Elements

Turn the page, throw a ball, pat your dog, rake leaves; everything we touch—from the water we drink to the air we breathe—is composed of matter. **Matter** refers to anything that takes up space and has mass. Although matter has many diverse forms—anything from molten lava to kidney stones—it only exists in three distinct states: solid, liquid, and gas.

### Elements

All matter, both nonliving and living, is composed of certain basic substances called **elements**. An element is a substance that cannot be broken down to simpler substances with different properties (a property is a physical or chemical characteristic, such as density, solubility, melting point, and reactivity) by ordinary chemical means. It is quite remarkable that there are only 92 naturally occurring elements that serve as the building blocks of matter. Other elements have been “human-made” and are not biologically important.

Both the Earth’s crust and all organisms are composed of elements, but they differ as to which ones are common. Only six elements—carbon, hydrogen, nitrogen, oxygen, phosphorus, and sulfur—are basic to life and make up about 95% of the body weight of organisms. The acronym CHNOPS helps us remember these six elements. The properties of these elements are essential to the uniqueness of cells and organisms, such as the macaws in Figure 2.1. The macaws have gathered on a salt lick in South America. Salt contains the elements sodium and chlorine and is commonly sought after by many forms of life. Potassium, calcium, iron, and magnesium are still other elements found in living things.

### Atoms

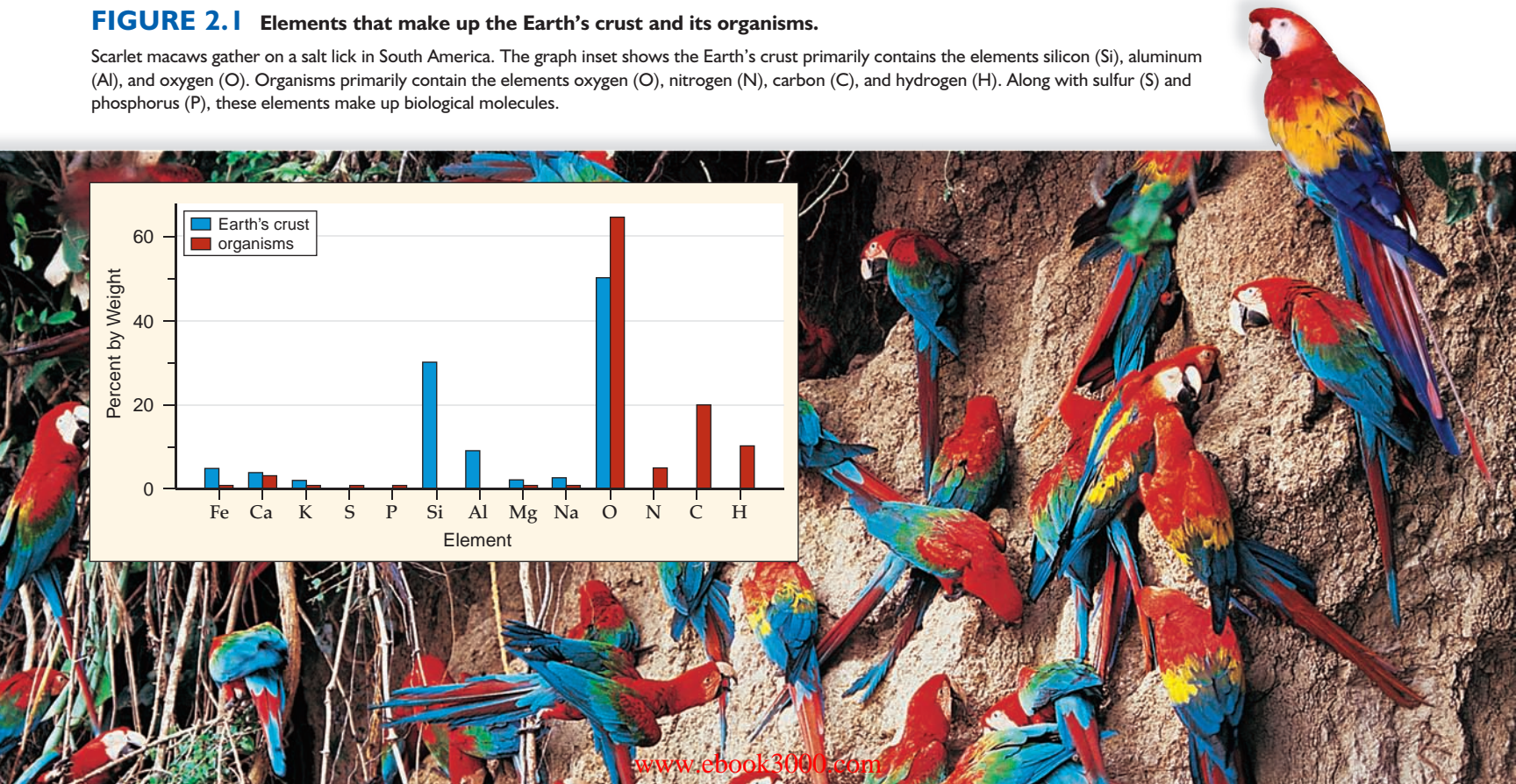
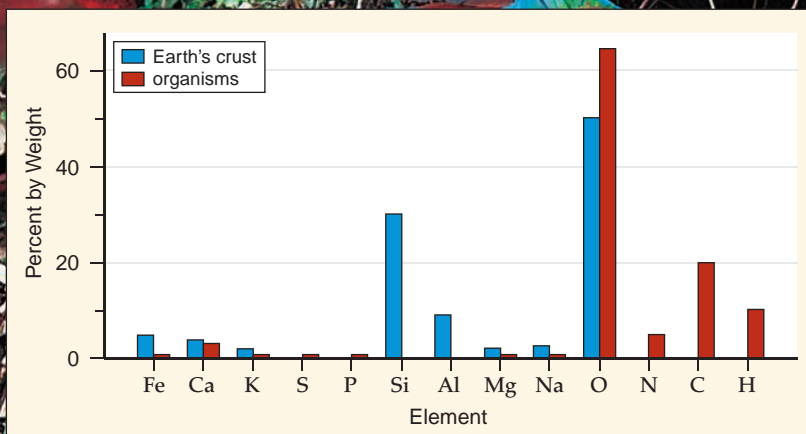
In the early 1800s, the English scientist John Dalton championed the atomic theory, which says that elements consist of tiny particles called **atoms** [Gk. *atomos*, uncut, indivisible]. An atom is the smallest part of an element that displays the properties of the element. An element and its atoms share the same name. One or two letters create the **atomic symbol**, which stands for this name. For example, the symbol H means a hydrogen atom, the symbol Rn stands for radon, and the symbol Na (for *natrium* in Latin) is used for a sodium atom.

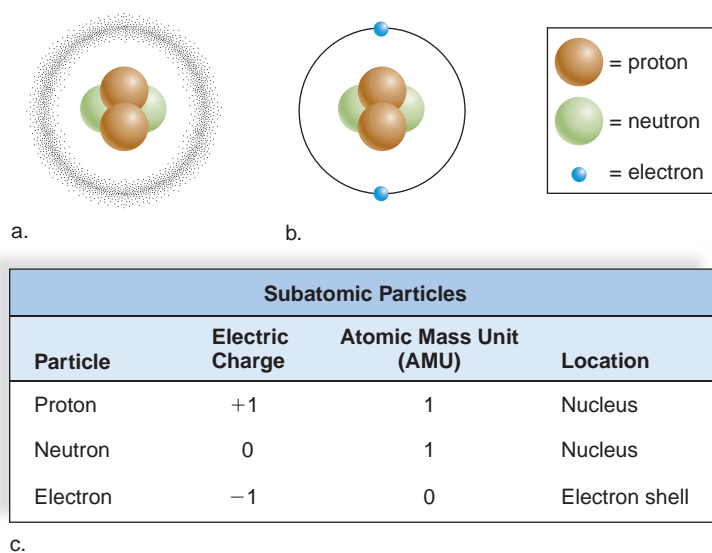
Physicists have identified a number of subatomic particles that make up atoms. The three best known subatomic particles include positively charged **protons**, uncharged **neutrons**, and negatively charged **electrons** [Gk. *elektron*, electricity]. Protons and neutrons are located within the nucleus of an atom, and electrons move about the nucleus. Figure 2.2 shows the arrangement of the subatomic particles in a helium atom, which has only two electrons. In Figure 2.2a, the stippling shows the probable location of electrons, and in Figure 2.2b, the circle represents an **electron shell**, the average location of electrons.

The concept of an atom has changed greatly since Dalton’s day. If an atom could be drawn the size of a football field, the nucleus would be like a gumball in the center of the field, and the electrons would be tiny specks whirling about in the upper stands. Most of an atom is empty space. We should also realize that we can only indicate where the electrons are expected to be most of the time. In our analogy, the electrons might very well stray outside the stadium at times.

**FIGURE 2.1** Elements that make up the Earth’s crust and its organisms.

Scarlet macaws gather on a salt lick in South America. The graph inset shows the Earth’s crust primarily contains the elements silicon (Si), aluminum (Al), and oxygen (O). Organisms primarily contain the elements oxygen (O), nitrogen (N), carbon (C), and hydrogen (H). Along with sulfur (S) and phosphorus (P), these elements make up biological molecules.



**FIGURE 2.2** Model of helium (He).

Atoms contain subatomic particles, which are located as shown. Protons and neutrons are found within the nucleus, and electrons are outside the nucleus. **a.** The stippling shows the probable location of the electrons in the helium atom. **b.** The average location of an electron is sometimes represented by a circle termed an electron shell. **c.** The electric charge and the atomic mass units (AMU) of the subatomic particles vary as shown.

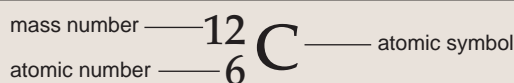
## Atomic Number and Mass Number

Atoms not only have an atomic symbol, they also have an atomic number and mass number. All atoms of an element have the same number of protons housed in the nucleus. This is called the **atomic number**, which accounts for the unique properties of this type atom.

Each atom also has its own mass number dependent on the number of subatomic particles in that atom. Protons and neutrons are assigned one atomic mass unit (AMU) each. Electrons are so small that their AMU is considered zero in most calculations (Fig. 2.2c). Therefore, the **mass number** of an atom is the sum of protons and neutrons in the nucleus.

The term mass is used, and not *weight*, because mass is constant, while weight changes according to the gravitational force of a body. The gravitational force of the Earth is greater than that of the moon; therefore, substances weigh less on the moon, even though their mass has not changed.

By convention, when an atom stands alone (and not in the periodic table, discussed next), the atomic number is written as a subscript to the lower left of the atomic symbol. The mass number is written as a superscript to the upper left of the atomic symbol. Regardless of position, the smaller number is always the atomic number, as shown here for carbon.



Periods	I							VIII
	1 <b>H</b> 1.008	atomic mass						2 <b>He</b> 4.003
		II	III	IV	V	VI	VII	
	3 <b>Li</b> 6.941	4 <b>Be</b> 9.012	5 <b>B</b> 10.81	6 <b>C</b> 12.01	7 <b>N</b> 14.01	8 <b>O</b> 16.00	9 <b>F</b> 19.00	10 <b>Ne</b> 20.18
	11 <b>Na</b> 22.99	12 <b>Mg</b> 24.31	13 <b>Al</b> 26.98	14 <b>Si</b> 28.09	15 <b>P</b> 30.97	16 <b>S</b> 32.07	17 <b>Cl</b> 35.45	18 <b>Ar</b> 39.95
	19 <b>K</b> 39.10	20 <b>Ca</b> 40.08	31 <b>Ga</b> 69.72	32 <b>Ge</b> 72.59	33 <b>As</b> 74.92	34 <b>Se</b> 78.96	35 <b>Br</b> 79.90	36 <b>Kr</b> 83.60
	Groups							

**FIGURE 2.3** A portion of the periodic table.

In the periodic table, the elements, and therefore atoms, are in the order of their atomic numbers but arranged so that they are placed in groups (vertical columns) and periods (horizontal rows). All the atoms in a particular group have chemical characteristics in common. These four periods contain the elements that are most important in biology; the complete periodic table is in Appendix D.

## The Periodic Table

Once chemists discovered a number of the elements, they began to realize that even though each element consists of a different atom, certain chemical and physical characteristics recur. The periodic table, developed by the Russian chemist Dmitri Mendeleev (1834–1907), was constructed as a way to group the elements, and therefore atoms, according to these characteristics.

Figure 2.3 is a portion of the periodic table, which is shown in total in Appendix D. The atoms shown in the periodic table are assumed to be electrically neutral. Therefore, the atomic number not only tells you the number of protons, it also tells you the number of electrons. The **atomic mass** is the average of the AMU for all the isotopes (discussed next) of that atom. To determine the number of neutrons, subtract the number of protons from the atomic mass, and take the closest whole number.

In the periodic table, every atom is in a particular period (the horizontal rows) and in a particular group (the vertical columns). The atomic number of every atom in a period increases by one if you read from left to right. All the atoms in a group share the same binding characteristics. For example, all the atoms in group VII react with one atom at a time, for reasons we will soon explore. The atoms in group VIII are called the noble gases because they are inert and rarely react with another atom. Notice that helium and krypton are noble gases.



## Isotopes

**Isotopes** [Gk. *isos*, equal, and *topos*, place] are atoms of the same element that differ in the number of neutrons. Isotopes have the same number of protons, but they have different atomic masses. For example, the element carbon has three common isotopes:



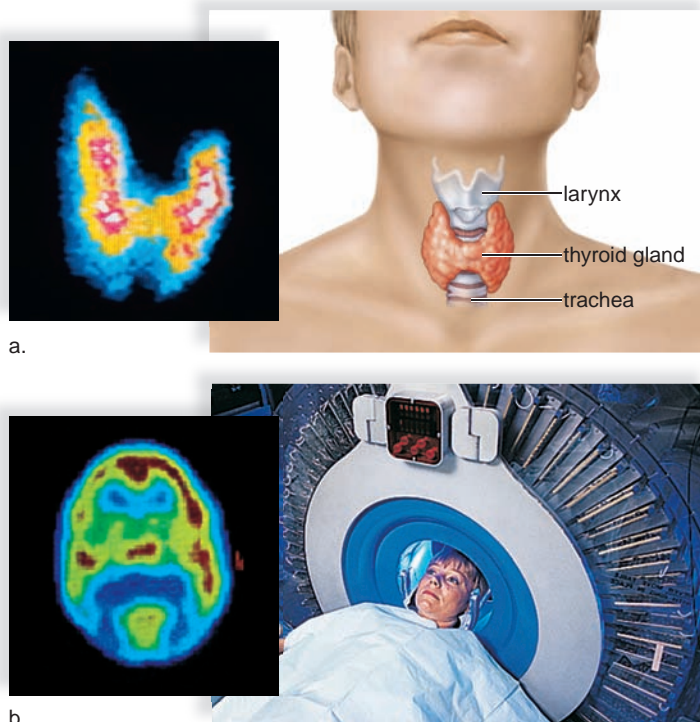
\*radioactive

Carbon 12 has six neutrons, carbon 13 has seven neutrons, and carbon 14 has eight neutrons. Unlike the other two isotopes of carbon, carbon 14 is unstable; it changes over time into nitrogen 14, which is a stable isotope of the element nitrogen. As carbon 14 decays, it releases various types of energy in the form of rays and subatomic particles, and therefore it is a radioactive isotope. The radiation given off by radioactive isotopes can be detected in various ways. The Geiger counter is an instrument that is commonly used to detect radiation. In 1860, the French physicist Antoine-Henri Becquerel discovered that a sample of uranium would produce a bright image on a photographic plate because it was radioactive. A similar method of detecting radiation is still in use today. Marie Curie, who worked with Becquerel, contributed much to the study of radioactivity, as she named it. Today, radiation is used by biologists to date objects, create images, and trace the movement of substances.

### Low Levels of Radiation

The chemical behavior of a radioactive isotope is essentially the same as that of the stable isotopes of an element. This means that you can put a small amount of radioactive isotope in a sample and it becomes a **tracer** by which to detect molecular changes. Melvin Calvin and his co-workers used carbon 14 to detect all the various reactions that occur during the process of photosynthesis.

The importance of chemistry to medicine is nowhere more evident than in the many medical uses of radioactive isotopes. Specific tracers are used in imaging the body's organs and tissues. For example, after a patient drinks a solution containing a minute amount of  ${}^{131}\text{I}$ , it becomes concentrated in the thyroid—the only organ to take it up. A subsequent image of the thyroid indicates whether it is healthy in structure and function (Fig. 2.4a). Positron-emission tomography (PET) is a way to determine the comparative activity of tissues. Radioactively labeled glucose, which emits a subatomic particle known as a positron, is injected into the body. The radiation given off is detected by sensors and analyzed by a computer. The result is a color image that shows which tissues took up glucose and are metabolically active. The red areas surrounded by green in Figure 2.4b indicate which areas of the brain are most active. PET scans of the brain are used to evaluate patients who have memory disorders of an undetermined cause or suspected brain tumors or seizure disorders that could possibly benefit from surgery. PET scans, utilizing radioactive thallium, can detect signs of coronary artery disease and low blood flow to the heart.

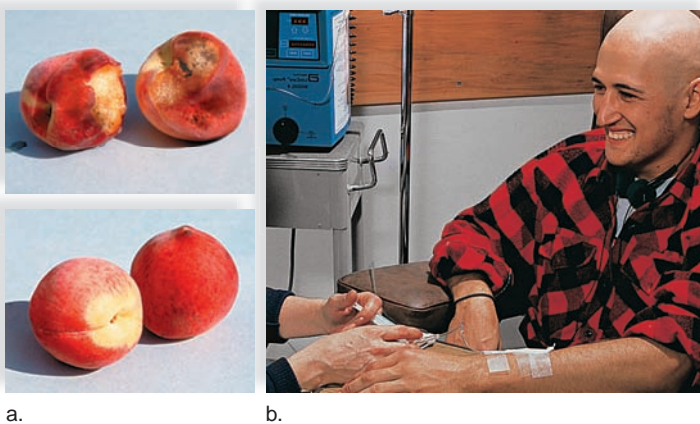


**FIGURE 2.4** Low levels of radiation.

**a.** Incomplete scan of the thyroid gland on the left indicates the presence of a tumor that does not take up the radioactive iodine. **b.** A PET (positron-emission tomography) scan reveals which portions of the brain are most active (green and red colors).

### High Levels of Radiation

Radioactive substances in the environment can harm cells, damage DNA, and cause cancer. When Marie Curie was studying radiation, its harmful effects were not known, and she and many of her co-workers developed cancer. The release of radioactive particles following a nuclear power plant accident can have far-reaching and long-lasting effects on human health. The harmful effects of radiation can be put to good use, however (Fig. 2.5). Radiation from radioactive isotopes has been used for many years to sterilize medical



**FIGURE 2.5** High levels of radiation.

**a.** Radiation kills bacteria and fungi. After irradiation, peaches spoil less quickly and can be kept for a longer length of time. **b.** Physicians use targeted radiation therapy to kill cancer cells.

and dental products. Now it can be used to sterilize the U.S. mail and other packages to free them of possible pathogens, such as anthrax spores. The ability of radiation to kill cells is often applied to cancer cells. Targeted radioisotopes can be introduced into the body so that the subatomic particles emitted destroy only cancer cells, with little risk to the rest of the body.

## Electrons and Energy

In an electrically neutral atom, the positive charges of the protons in the nucleus are balanced by the negative charges of electrons moving about the nucleus. Various models in years past have attempted to illustrate the precise location of electrons. Figure 2.6 uses the Bohr model, which is named after the physicist Niels Bohr. The Bohr model is useful, but we need to realize that today's physicists tell us it is not possible to determine the precise location of any individual electron at any given moment.

In the Bohr model, the electron shells about the nucleus also represent energy levels. It seems reasonable to suggest that negatively charged electrons are attracted to the positively charged nucleus, and that it takes energy to push them away and keep them in their own shell. Further, the more distant the shell, the more energy it takes. Therefore, it is proper to speak of electrons as being at particular energy levels in relation to the nucleus. When you study photosynthesis, you will learn that when atoms absorb the energy of the sun, electrons are boosted to a higher energy level. Later, as the electrons return to their original energy level, energy is released and transformed into chemical energy. This chemical energy supports all life on Earth and therefore our very existence is dependent on the energy of electrons.

You will want to learn to draw a Bohr model for each of the elements that occurs in the periodic table shown in Figure 2.3. Let's begin by examining the models depicted in Figure 2.6. Notice that the first shell (closest to the nucleus) can contain two electrons; thereafter, each additional shell can

contain eight electrons. Also, each lower level is filled with electrons before the next higher level contains any electrons.

The sulfur atom, with an atomic number of 16, has two electrons in the first shell, eight electrons in the second shell, and six electrons in the third, or outer, shell. Revisit the periodic table (see Fig. 2.3), and note that sulfur is in the third period. In other words, the period tells you how many shells an atom has. Also note that sulfur is in group VI. The group tells you how many electrons an atom has in its outer shell.

If an atom has only one shell, the outer shell is complete when it has two electrons. Otherwise, the **octet rule**, which states that the outer shell is most stable when it has eight electrons, holds. As mentioned previously, atoms in group VIII of the periodic table are called the noble gases because they do not ordinarily react. Stability exists because an outer shell with eight electrons has less energy. In general, lower energy states represent stability, as we will have an opportunity to point out again in Chapter 6.

Just as you sometimes communicate with and react to other people by using your hands, so atoms use the electrons in their outer shells to undergo reactions. Atoms with fewer than eight electrons in the outer shell react with other atoms in such a way that after the reaction, each has a stable outer shell. As we shall see, the number of electrons in an atom's outer shell, called the **valence shell**, determines whether it gives up, accepts, or shares electrons to acquire eight electrons in the outer shell.

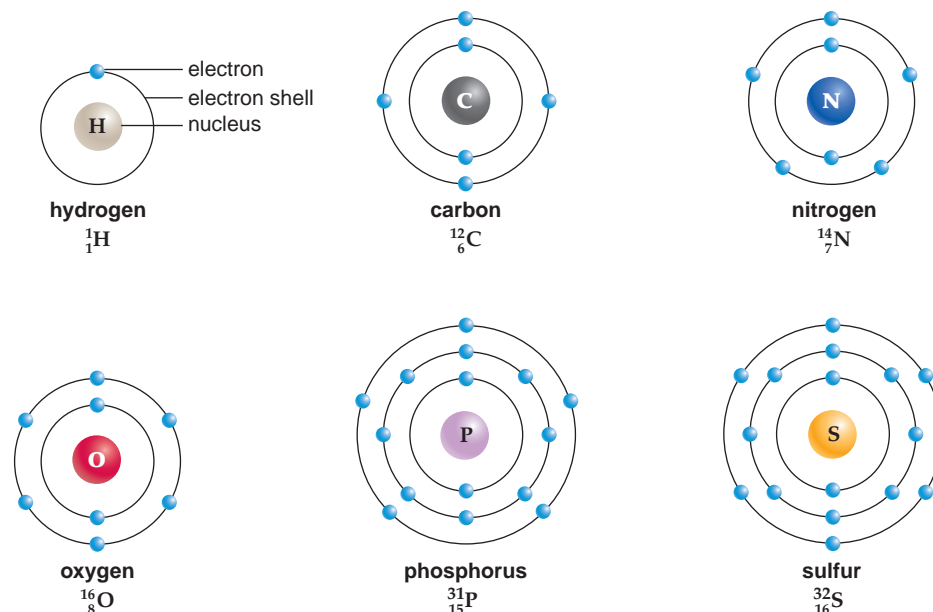
### Check Your Progress

### 2.1

1. Contrast atomic number and mass number.
2. **a.** How do group III elements differ in the periodic table? **b.** How do period III elements differ?
3. List some uses of radioactive isotopes in biology and medicine.

**FIGURE 2.6** Bohr models of atoms.

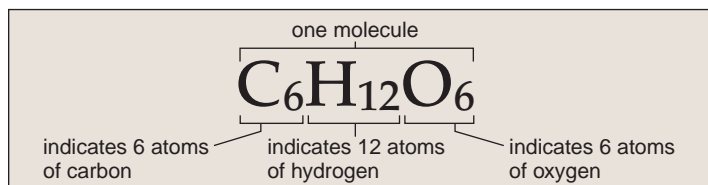
Electrons orbit the nucleus at particular energy levels (electron shells): The first shell contains up to two electrons, and each shell thereafter can contain up to eight electrons as long as we consider only atoms with an atomic number of 20 or below. Each shell is filled before electrons are placed in the next shell. Why does carbon have only two shells while phosphorus and sulfur have three shells?



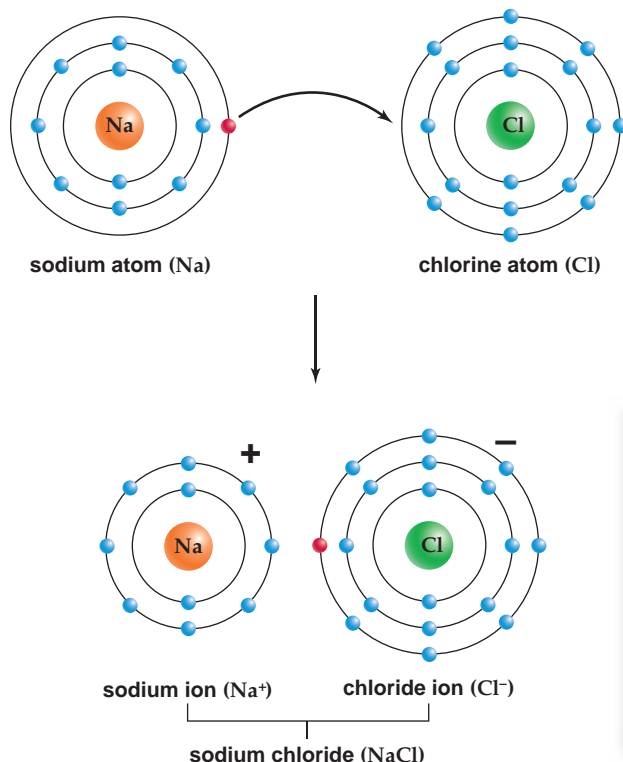


## 2.2 Compounds and Molecules

A **compound** exists when two or more elements have bonded together. A **molecule** [L. *moles*, mass] is the smallest part of a compound that still has the properties of the particular compound. In practice, these two terms are used interchangeably, but in biology, we usually speak of molecules. Water ( $\text{H}_2\text{O}$ ) is a molecule that contains atoms of hydrogen and oxygen. A **formula** tells you the number of each kind of atom in a molecule. For example, in glucose:



Electrons possess energy, and the bonds that exist between atoms also contain energy. Organisms are directly dependent on chemical-bond energy to maintain their organization. As you may know, organisms routinely break down glucose, the sugar shown above, to obtain energy. When a chemical reaction occurs, as when glucose is broken down, electrons shift in their relationship to one another, and energy is released. Spontaneous reactions, which are ones that occur freely, always release energy.



a.

### Ionic Bonding

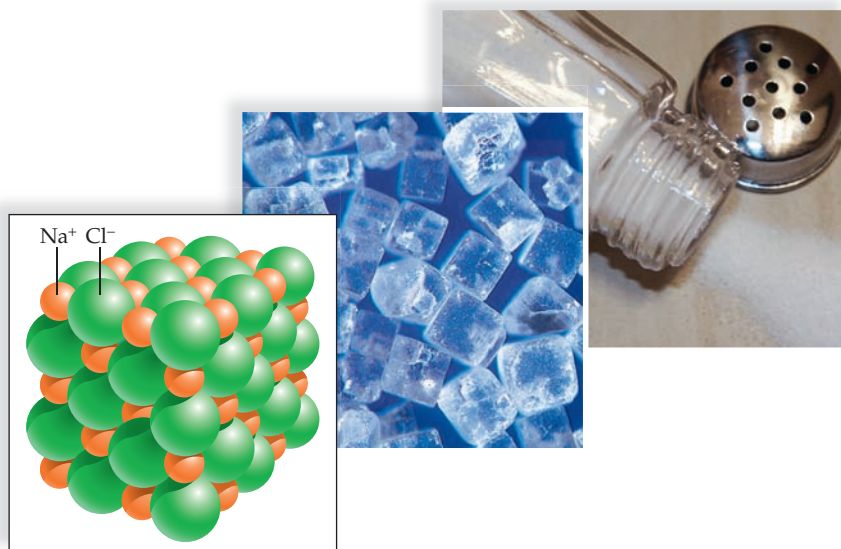
Sodium (Na), with only one electron in its third shell, tends to be an electron donor (Fig. 2.7a). Once it gives up this electron, the second shell, with eight electrons, becomes its outer shell. Chlorine (Cl), on the other hand, tends to be an electron acceptor. Its outer shell has seven electrons, so if it acquires only one more electron it has a completed outer shell. When a sodium atom and a chlorine atom come together, an electron is transferred from the sodium atom to the chlorine atom. Now both atoms have eight electrons in their outer shells.

This electron transfer, however, causes a charge imbalance in each atom. The sodium atom has one more proton than it has electrons; therefore, it has a net charge of +1 (symbolized by  $\text{Na}^+$ ). The chlorine atom has one more electron than it has protons; therefore, it has a net charge of -1 (symbolized by  $\text{Cl}^-$ ). Such charged particles are called **ions**. Sodium ( $\text{Na}^+$ ) and chloride ( $\text{Cl}^-$ ) are not the only biologically important ions. Some, such as potassium ( $\text{K}^+$ ), are formed by the transfer of a single electron to another atom; others, such as calcium ( $\text{Ca}^{2+}$ ) and magnesium ( $\text{Mg}^{2+}$ ), are formed by the transfer of two electrons.

Ionic compounds are held together by an attraction between negatively and positively charged ions called an **ionic bond**. When sodium reacts with chlorine, an ionic compound called sodium chloride ( $\text{NaCl}$ ) results. Sodium chloride is a salt, commonly known as table salt, because it is used to season our food (Fig. 2.7b). **Salts** are solid substances that usually separate and exist as individual ions in water, as discussed on page 30.

**FIGURE 2.7** Formation of sodium chloride (table salt).

**a.** During the formation of sodium chloride, an electron is transferred from the sodium atom to the chlorine atom. At the completion of the reaction, each atom has eight electrons in the outer shell, but each also carries a charge as shown. **b.** In a sodium chloride crystal, ionic bonding between  $\text{Na}^+$  and  $\text{Cl}^-$  causes the atoms to assume a three-dimensional lattice in which each sodium ion is surrounded by six chloride ions, and each chloride ion is surrounded by six sodium ions. The result is crystals of salt as in table salt.



b.

## Covalent Bonding

A **covalent bond** [L. *co*, together, with, and *valens*, strength] results when two atoms share electrons in such a way that each atom has an octet of electrons in the outer shell (or two electrons, in the case of hydrogen). In a hydrogen atom, the outer shell is complete when it contains two electrons. If hydrogen is in the presence of a strong electron acceptor, it gives up its electron to become a hydrogen ion ( $\text{H}^+$ ). But if this is not possible, hydrogen can share with another atom and thereby have a completed outer shell. For example, one hydrogen atom will share with another hydrogen atom. Their two electron shells overlap and the electrons are shared between them (Fig. 2.8a). Because they share the electron pair, each atom has a completed outer shell.

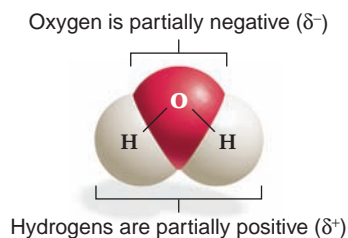
A more common way to symbolize that atoms are sharing electrons is to draw a line between the two atoms, as in the structural formula  $\text{H}-\text{H}$ . Just as a handshake requires two hands, one from each person, a covalent bond between two atoms requires two electrons, one from each atom. In a molecular formula, the line is omitted and the molecule is simply written as  $\text{H}_2$ .

Sometimes, atoms share more than one pair of electrons to complete their octets. A double covalent bond occurs when two atoms share two pairs of electrons (Fig. 2.8b). To show that oxygen gas ( $\text{O}_2$ ) contains a double bond, the molecule can be written as  $\text{O}=\text{O}$ . It is also possible for atoms to form triple covalent bonds, as in nitrogen gas ( $\text{N}_2$ ), which can be written as  $\text{N}\equiv\text{N}$ . Single covalent bonds between atoms are quite strong, but double and triple bonds are even stronger.

### Nonpolar and Polar Covalent Bonds

When the sharing of electrons between two atoms is equal, the covalent bond is said to be a **nonpolar covalent bond**. If one atom is able to attract electrons to a greater degree than the other atom, it is the more electronegative atom. **Electronegativity** is dependent on the number of protons—the greater the number of protons, the greater the electronegativity. When electrons are not shared equally, the covalent bond is a **polar covalent bond**.

You can readily see that the bonds in methane (Fig. 2.8c) must be polar because carbon has more protons than a hydrogen atom. However, methane is a symmetrical molecule and the polarities cancel each other out—methane is a nonpolar molecule. Not so in water, which has this shape:



In water, the oxygen atom is more electronegative than the hydrogen atoms and the bonds are polar. Moreover, because of its nonsymmetrical shape, the polar bonds cannot cancel

Electron Model	Structural Formula	Molecular Formula
	$\text{H}-\text{H}$	$\text{H}_2$
a. Hydrogen gas		
	$\text{O}=\text{O}$	$\text{O}_2$
b. Oxygen gas		
	$\begin{array}{c} \text{H} \\   \\ \text{H}-\text{C}-\text{H} \\   \\ \text{H} \end{array}$	$\text{CH}_4$
c. Methane		

**FIGURE 2.8** Covalently bonded molecules.

In a covalent bond, atoms share electrons, allowing each atom to have a completed outer shell. **a.** A molecule of hydrogen ( $\text{H}_2$ ) contains two hydrogen atoms sharing a pair of electrons. This single covalent bond can be represented in any of the three ways shown. **b.** A molecule of oxygen ( $\text{O}_2$ ) contains two oxygen atoms sharing two pairs of electrons. This results in a double covalent bond. **c.** A molecule of methane ( $\text{CH}_4$ ) contains one carbon atom bonded to four hydrogen atoms.

each other and water is a polar molecule. The more electronegative end of the molecule is designated slightly negative ( $\delta^-$ ), and the hydrogens are designated slightly positive ( $\delta^+$ ).

Water is not the only polar molecule in living things. For example, the amine group ( $-\text{NH}_2$ ) is polar, and this causes amino acids and nucleic acids to exhibit polarity, as we shall see in the next chapter. The polarity of molecules affects how they interact with other molecules.

### Check Your Progress

### 2.2

1. Contrast an ionic bond with a covalent bond.
2. Why would you expect calcium to become an ion that carries two plus charges?
3. Explain how it is that all the atoms in methane ( $\text{CH}_4$ ) have a complete outer shell.



## 2.3 Chemistry of Water

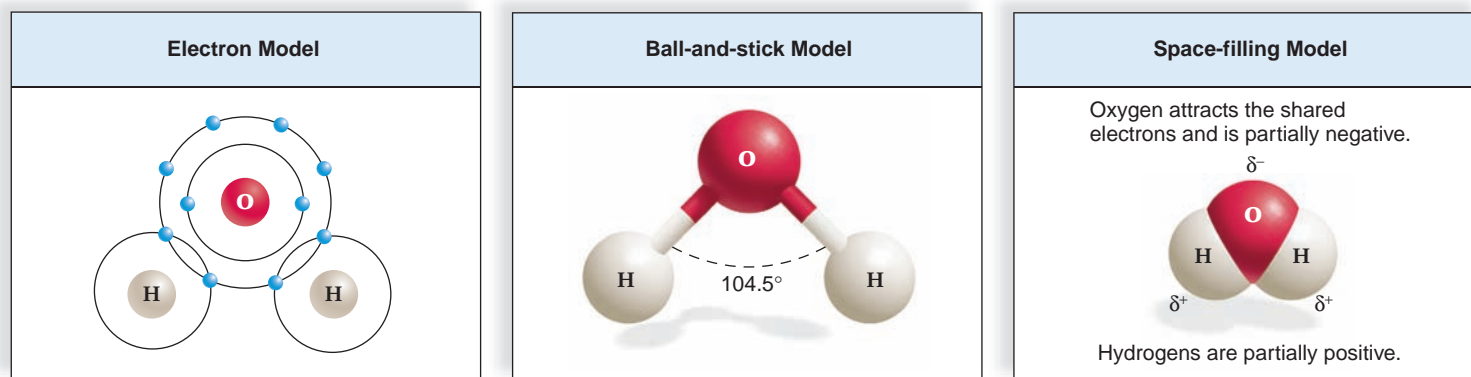
Figure 2.9a recaps what we know about the water molecule. The structural formula on the far left shows that when water forms, an oxygen atom is sharing electrons with two hydrogen atoms. The ball-and-stick model in the center shows that the covalent bonds between oxygen and each of the hydrogens are at an angle of  $104.5^\circ$ . Finally, the space-filling molecule gives us the three-dimensional shape of the molecule and indicates its polarity.

The shape of water and of all organic molecules is necessary to the structural and functional roles they play in living things. For example, hormones have specific shapes that allow them to be recognized by the cells in the body. We can stay well only when antibodies combine with disease-causing agents, like a key fits a lock. Similarly, homeostasis is only maintained when enzymes have the proper shape to carry out their particular reactions in cells. The shape of a water molecule and its polarity makes hydrogen bonding possible. A **hydrogen bond** is the attraction of a slightly positive hydrogen to a slightly negative atom in the vicinity. In carbon dioxide,  $\text{O}=\text{C}=\text{O}$ , there is also a slight difference in polarity between carbon and the oxygens but because carbon dioxide is symmetrical, the opposing charges cancel one another and hydrogen bonding does not occur.

## Hydrogen Bonding

The dotted lines in Figure 2.9b indicate that the hydrogen atoms in one water molecule are attracted to the oxygen atoms in other water molecules. This attraction, which is weaker than an ionic or covalent bond, is called a hydrogen bond. The dotted lines indicate that hydrogen bonds are more easily broken than covalent bonds. Hydrogen bonding is not unique to water. Other biological molecules, such as DNA, have polar covalent bonds involving an electropositive hydrogen and usually an electronegative oxygen or nitrogen. In these instances, a hydrogen bond can occur within the same molecule or between nearby molecules.

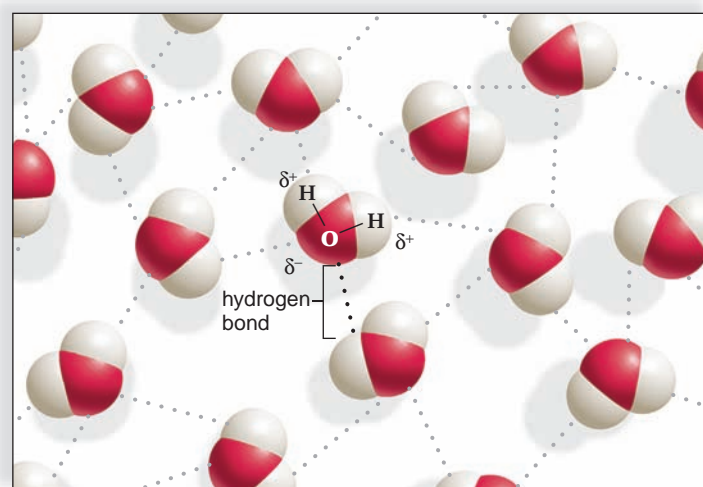
Although a hydrogen bond is more easily broken than a covalent bond, many hydrogen bonds taken together are quite strong. Hydrogen bonds between cellular molecules help maintain their proper structure and function. For example, hydrogen bonds hold the two strands of DNA together. When DNA makes a copy of itself, hydrogen bonds easily break, allowing DNA to unzip. But normally, the hydrogen bonds add stability to the DNA molecule. Similarly, the shape of protein molecules is often maintained by hydrogen bonding between parts of the same molecule. As we shall see, many of the important properties of water are the result of hydrogen bonding.



a. Water ( $\text{H}_2\text{O}$ )

**FIGURE 2.9** Water molecule.

**a.** Three models for the structure of water. The electron model does not indicate the shape of the molecule. The ball-and-stick model shows that the two bonds in a water molecule are angled at  $104.5^\circ$ . The space-filling model also shows the V shape of a water molecule. **b.** Hydrogen bonding between water molecules. Each water molecule can hydrogen-bond to four other molecules. When water is in its liquid state, some hydrogen bonds are forming and others are breaking at all times.



b. Hydrogen bonding between water molecules

## Properties of Water

The first cell(s) evolved in water, and all living things are 70–90% water. Due to hydrogen bonding, water molecules cling together. Without hydrogen bonding between molecules, water would melt at  $-100^{\circ}\text{C}$  and boil at  $-91^{\circ}\text{C}$ , making most of the water on Earth steam, and life unlikely. But because of hydrogen bonding, water is a liquid at temperatures typically found on the Earth's surface. It melts at  $0^{\circ}\text{C}$  and boils at  $100^{\circ}\text{C}$ . These and other unique properties of water make it essential to the existence of life as we know it. When scientists examine the other planets with the hope of finding life, they first look for signs of water.

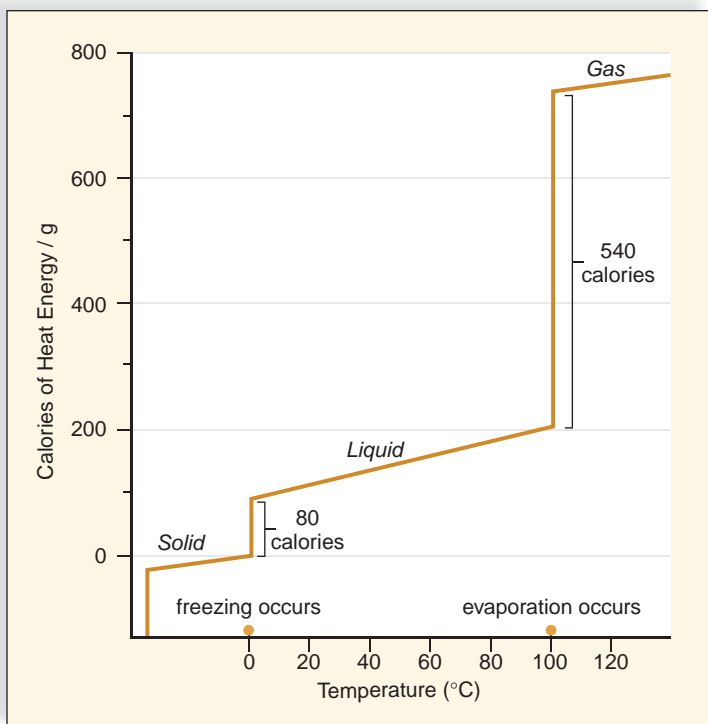
**Water Has a High Heat Capacity.** A **calorie** is the amount of heat energy needed to raise the temperature of 1 g of water  $1^{\circ}\text{C}$ . In comparison, other covalently bonded liquids require input of only about half this amount of energy to rise in temperature  $1^{\circ}\text{C}$ . The many hydrogen bonds that link water molecules together help water absorb heat without a great change in temperature. Converting 1 g of the coldest liquid water to ice requires the loss of 80 calories of heat energy (Fig. 2.10a). Water holds onto its heat, and its temperature falls more slowly than that of other liquids. This property of water is important not only for aquatic organisms but also for all living things.

Because the temperature of water rises and falls slowly, organisms are better able to maintain their normal internal temperatures and are protected from rapid temperature changes.

**Water Has a High Heat of Evaporation.** When water boils, it evaporates—that is, vaporizes into the environment. Converting 1 g of the hottest water to a gas requires an input of 540 calories of energy. Water has a high heat of evaporation because hydrogen bonds must be broken before water boils.

Water's high heat of vaporization gives animals in a hot environment an efficient way to release excess body heat. When an animal sweats, or gets splashed, body heat is used to vaporize water, thus cooling the animal (Fig. 2.10b). Because of water's high heat of vaporization and ability to hold onto its heat, temperatures along the coasts are moderate. During the summer, the ocean absorbs and stores solar heat, and during the winter, the ocean releases it slowly. In contrast, the interior regions of continents experience abrupt changes in temperatures.

**Water Is a Solvent.** Due to its polarity, water facilitates chemical reactions, both outside and within living systems. It dissolves a great number of substances. A **solution** contains dissolved substances, which are then called **solutes**.



a. Calories lost when 1 g of liquid water freezes and calories required when 1 g of liquid water evaporates.



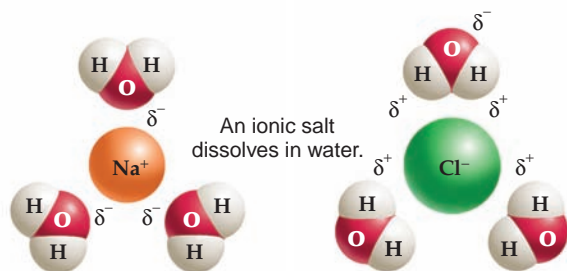
b. Bodies of organisms cool when their heat is used to evaporate water.

### FIGURE 2.10 Temperature and water.

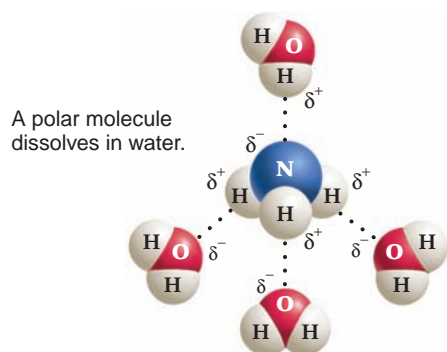
**a.** Water can be a solid, a liquid, or a gas at naturally occurring environmental temperatures. At room temperature and pressure, water is a liquid. When water freezes and becomes a solid (ice), it gives off heat, and this heat can help keep the environmental temperature higher than expected. On the other hand, when water evaporates, it takes up a large amount of heat as it changes from a liquid to a gas. **b.** This means that splashing water on the body will help keep body temperature within a normal range. Can you also see why water's properties help keep the coasts moderate in both winter and summer?



When ionic salts—for example, sodium chloride (NaCl)—are put into water, the negative ends of the water molecules are attracted to the sodium ions, and the positive ends of the water molecules are attracted to the chloride ions. This causes the sodium ions and the chloride ions to separate, or dissociate, in water.



Water is also a solvent for larger polar molecules, such as ammonia (NH<sub>3</sub>).



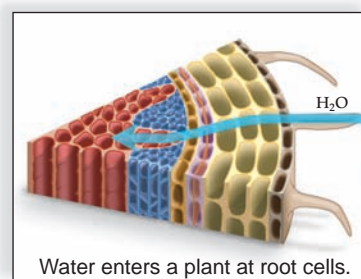
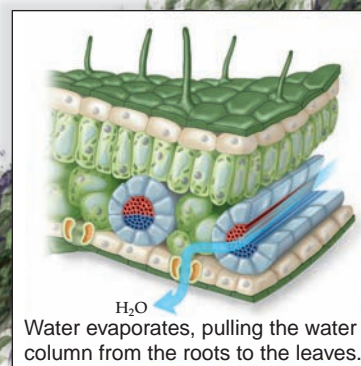
Those molecules that can attract water are said to be **hydrophilic** [Gk. *hydrias*, of water, and *phileo*, love]. When ions and molecules disperse in water, they move about and collide, allowing reactions to occur. Nonionized and nonpolar molecules that cannot attract water are said to be **hydrophobic** [Gk. *hydrias*, of water, and *phobos*, fear]. Gasoline contains nonpolar molecules, and therefore it does not mix with water and is hydrophobic.

**Water Molecules Are Cohesive and Adhesive.** Cohesion refers to the ability of water molecules to cling to each other due to hydrogen bonding. Because of cohesion, water exists as a liquid under ordinary conditions of temperature and pressure. The strong cohesion of water molecules is apparent because water flows freely, yet water molecules do not separate from each other. Adhesion refers to the ability of water molecules to cling to other polar surfaces. This is because of water's polarity. Multicellular animals often contain internal vessels in which water assists the transport of nutrients and wastes because the cohesion and adhesion of water allows blood to fill the tubular vessels of the cardiovascular system. For example, the liquid portion of our blood, which transports dissolved and suspended substances about the body, is 90% water.

Cohesion and adhesion also contribute to the transport of water in plants. Plants have their roots anchored in the soil, where they absorb water, but the leaves are uplifted

and exposed to solar energy. Water evaporating from the leaves is immediately replaced with water molecules from transport vessels that extend from the roots to the leaves (Fig. 2.11). Because water molecules are cohesive, a tension is created that pulls the water column up from the roots. Adhesion of water to the walls of the vessels also helps prevent the water column from breaking apart.

Because water molecules are strongly attracted to each other, they cling together at a surface exposed to air. The stronger the force between molecules in a liquid, the greater the **surface tension**. Water's high surface tension makes it possible for humans to skip rocks on water. Water striders, a common insect, can even walk on the surface of a pond without breaking the surface.



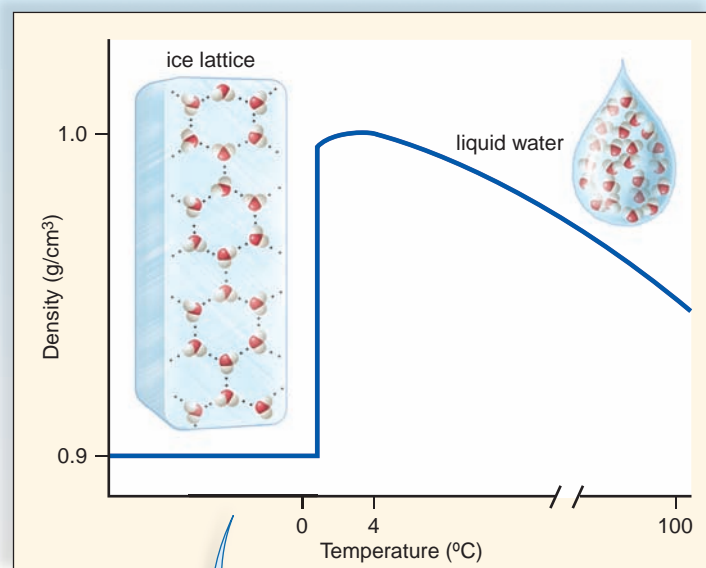
**FIGURE 2.11 Water as a transport medium.**

How does water rise to the top of tall trees? Vessels are water-filled pipelines from the roots to the leaves. When water evaporates from the leaves, the water column is pulled upward due to the cohesion of water molecules with one another and the adhesion of water molecules to the sides of the vessels.



**Frozen water (ice) is less dense than liquid water.** As liquid water cools, the molecules come closer together. Water is most dense at 4°C, but the water molecules are still moving about (Fig. 2.12). At temperatures below 4°C, there is only vibrational movement, and hydrogen bonding becomes more rigid but also more open. This means that water expands as it freezes, which is why cans of soda burst when placed in a freezer or why frost heaves make northern roads bumpy in the winter. It also means that ice is less dense than liquid water, and therefore ice floats on liquid water.

If ice did not float on water, it would sink, and ponds, lakes, and perhaps even the ocean would freeze solid, making life impossible in the water and also on land. Instead, bodies of water always freeze from the top down. When a body of water freezes on the surface, the ice acts as an insulator to prevent the water below it from freezing. This protects aquatic organisms so that they can survive the winter. As ice melts in the spring, it draws heat from the environment, helping to prevent a sudden change in temperature that might be harmful to life.



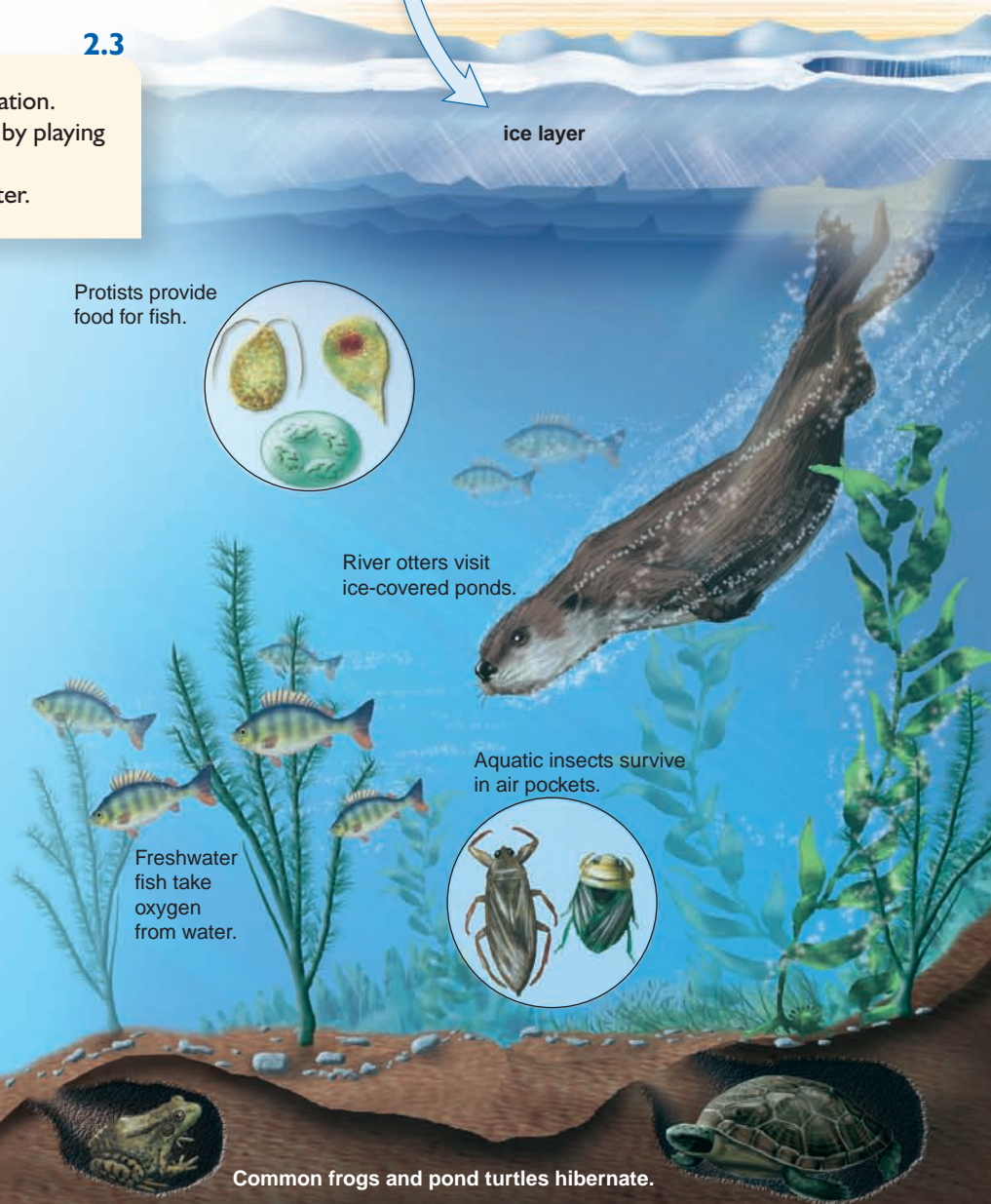
### Check Your Progress

### 2.3

1. Explain why water has a high heat of vaporization.
2. Explain why children in summer can cool off by playing in a sprinkler.
3. Explain why ice skating is possible in the winter.

**FIGURE 2.12 A pond in winter.**

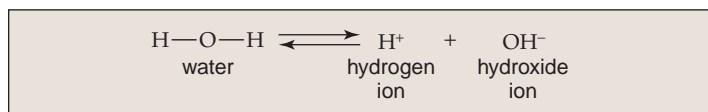
Above: Remarkably, water is more dense at 4°C than at 0°C. Most substances contract when they solidify, but water expands when it freezes because in ice, water molecules form a lattice in which the hydrogen bonds are farther apart than in liquid water. Below: The layer of ice that forms at the top of a pond shields the water and protects the protists, plants, and animals so that they can survive the winter. These animals, except for the otter, are ectothermic, which means that they take on the temperature of the outside environment. This might seem disadvantageous until you realize that water remains relatively warm because of its high heat capacity. During the winter, frogs and turtles hibernate and in this way, lower their oxygen needs. Insects survive in air pockets. Fish, as you will learn later in this text, have an efficient means of extracting oxygen from the water and they need less oxygen than the endothermic otter, which depends on muscle activity to warm its body.





## 2.4 Acids and Bases

When water ionizes, it releases an equal number of **hydrogen ions** ( $H^+$ ) (also called a proton<sup>1</sup>) and **hydroxide ions** ( $OH^-$ ):



Only a few water molecules at a time dissociate, and the actual number of  $H^+$  and  $OH^-$  is very small ( $1 \times 10^{-7}$  moles/liter).<sup>2</sup>

### Acidic Solutions (High $H^+$ Concentrations)

Lemon juice, vinegar, tomatoes, and coffee are all acidic solutions. What do they have in common? **Acids** are substances that dissociate in water, releasing hydrogen ions ( $H^+$ ). The acidity of a substance depends on how fully it dissociates in water. For example, hydrochloric acid (HCl) is a strong acid that dissociates almost completely in this manner:



If hydrochloric acid is added to a beaker of water, the number of hydrogen ions ( $H^+$ ) increases greatly.

### Basic Solutions (Low $H^+$ Concentration)

Milk of magnesia and ammonia are common basic solutions familiar to most people. **Bases** are substances that either take up hydrogen ions ( $H^+$ ) or release hydroxide ions ( $OH^-$ ). For example, sodium hydroxide (NaOH) is a strong base that dissociates almost completely in this manner:



If sodium hydroxide is added to a beaker of water, the number of hydroxide ions increases.

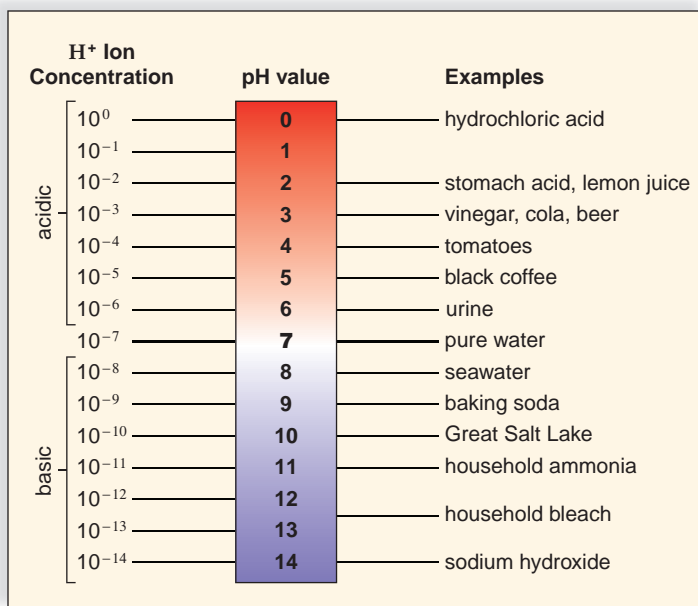
### pH Scale

The **pH scale** is used to indicate the acidity or basicity (alkalinity) of a solution.<sup>3</sup> The pH scale (Fig. 2.13) ranges from 0 to 14. A pH of 7 represents a neutral state in which the hydrogen ion and hydroxide ion concentrations are equal. A pH below 7 is an acidic solution because the hydrogen ion concentration is greater than the hydroxide concentration. A pH above 7 is basic because the  $[OH^-]$  is greater than the  $[H^+]$ . Further, as we move down the pH scale from pH 14 to pH 0, each unit is 10 times more acidic than the previous unit. As we move up the scale from 0 to 14, each unit is 10 times more basic than the previous unit. Therefore pH 5 is 100 times more acidic than is pH 7 and a 100 times more basic than pH 3.

<sup>1</sup> A hydrogen atom contains one electron and one proton. A hydrogen ion has only one proton, so it is often simply called a proton.

<sup>2</sup> In chemistry, a mole is defined as the amount of matter that contains as many objects (atoms, molecules, ions) as the number of atoms in exactly 12 g of  $^{12}\text{C}$ .

<sup>3</sup> pH is defined as the negative log of the hydrogen ion concentration  $[H^+]$ . A log is the power to which 10 must be raised to produce a given number.



**FIGURE 2.13** The pH scale.

The pH scale ranges from 0 to 14 with 0 being the most acidic and 14 being the most basic. pH 7 (neutral pH) has equal amounts of hydrogen ions ( $H^+$ ) and hydroxide ions ( $OH^-$ ). An acidic pH has more  $H^+$  than  $OH^-$  and a basic pH has more  $OH^-$  than  $H^+$ .

The pH scale was devised to eliminate the use of cumbersome numbers. For example, the possible hydrogen ion concentrations of a solution are on the left of this listing and the pH is on the right:

	$[H^+]$ (moles per liter)	pH
0.000001	$= 1 \times 10^{-6}$	6
0.0000001	$= 1 \times 10^{-7}$	7
0.00000001	$= 1 \times 10^{-8}$	8

To further illustrate the relationship between hydrogen ion concentration and pH, consider the following question. Which of the pH values listed indicates a higher hydrogen ion concentration  $[H^+]$  than pH 7, and therefore would be an acidic solution? A number with a smaller negative exponent indicates a greater quantity of hydrogen ions than one with a larger negative exponent. Therefore, pH 6 is an acidic solution.

The Ecology Focus on page 33 describes detrimental environmental consequences to nonliving and living things as rain and snow have become more acidic. In humans, pH needs to be maintained with a narrow range or there are health consequences. The pH of blood is around 7.4, and blood is buffered in the manner described next to keep the pH within a normal range.

### Buffers and pH

A **buffer** is a chemical or a combination of chemicals that keeps pH within normal limits. Many commercial products such as Bufferin, shampoos, or deodorants are buffered as an added incentive for us to buy them.

In living things, the pH of body fluids is maintained within a narrow range, or else health suffers. The pH of our blood when we are healthy is always about 7.4—that is, just

## The Harm Done by Acid Deposition

**N**ormally, rainwater has a pH of about 5.6 because the carbon dioxide in the air combines with water to give a weak solution of carbonic acid. Acid deposition includes rain or snow that has a pH of less than 5, as well as dry acidic particles that fall to Earth from the atmosphere. When fossil fuels such as coal, oil, and gasoline are burned, sulfur dioxide and nitrogen oxides combine with water to produce sulfuric and nitric acids. These pollutants are generally found eastward of where they originated because of wind patterns. The use of very tall smokestacks causes them to be carried even hundreds of miles away. For example, acid rain in southeastern Canada results from the burning of fossil fuels in factories and power plants in the midwestern United States.

### Impact on Lakes

Acid rain adversely affects lakes, particularly in areas where the soil is thin and lacks limestone (calcium carbonate, or  $\text{CaCO}_3$ ), a buffer to acid deposition. Acid deposition leaches toxic aluminum from the soil and converts mercury deposits in lake bottom sediments to toxic methyl mercury, which accumulates in fish. People are now advised against eating fish from the Great Lakes because of high mercury levels. Hundreds of lakes are devoid of fish in Canada and New England, and thousands have suffered the same

fate in the Scandinavian countries. Some of these lakes have no signs of life at all.

### Impact on Forests

The leaves of plants damaged by acid rain can no longer carry on photosynthesis as before. When plants are under stress, they become susceptible to diseases and pests of all types. Forests on mountaintops receive more rain than those at lower levels; therefore, they are more affected by acid rain (Fig. 2Aa). Forests are also damaged when toxic chemicals such as aluminum are leached from the soil. These kill soil fungi



a.

that assist roots in acquiring the nutrients trees need. In New England, 1.3 million acres of high-elevation forests have been devastated.

### Impact on Humans and Structures

Humans may be affected by acid rain. Inhaling dry sulfate and nitrate particles appears to increase the occurrence of respiratory illnesses, such as asthma. Buildings and monuments made of limestone and marble break down when exposed to acid rain (Fig. 2Ab). The paint on homes and automobiles is likewise degraded.



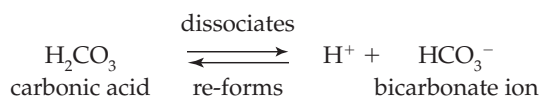
b.

**FIGURE 2A Effects of acid deposition.**

The burning of gasoline derived from oil, a fossil fuel, leads to acid deposition, which causes (a) trees to die and (b) statues to deteriorate.

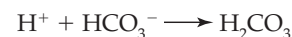
slightly basic (alkaline). If the blood pH drops to about 7, acidosis results. If the blood pH rises to about 7.8, alkalosis results. Both conditions can be life threatening; the blood pH must be kept around 7.4. Normally, pH stability is possible because the body has built-in mechanisms to prevent pH changes. Buffers are one of these important mechanisms.

Buffers help keep the pH within normal limits because they are chemicals or combinations of chemicals that take up excess hydrogen ions ( $\text{H}^+$ ) or hydroxide ions ( $\text{OH}^-$ ). For example, carbonic acid ( $\text{H}_2\text{CO}_3$ ) is a weak acid that minimally dissociates and then re-forms in the following manner:



Blood always contains a combination of some carbonic acid and some bicarbonate ions. When hydrogen ions

( $\text{H}^+$ ) are added to blood, the following reaction reduces acidity:



When hydroxide ions ( $\text{OH}^-$ ) are added to blood, this reaction reduces basidity:



These reactions prevent any significant change in blood pH.

### Check Your Progress

2.4

1. Contrast an acid with a base.
2. Give an example to substantiate that acid rain is detrimental to both plants and animals.
3. A substance that absorbs hydrogen ions makes the pH rise. Explain.



## Connecting the Concepts

All matter consists of various combinations of the same 92 elements. Living things consist primarily of just six of these elements—carbon, hydrogen, nitrogen, oxygen, phosphorus, and sulfur (CHNOPS for short). These elements combine to form the unique types of molecules found in living cells. In organisms, many other elements exist in smaller amounts as ions, and their functions are dependent on their charged nature. Cells consist largely of water, a molecule that contains only hydrogen and oxygen. Polar covalent bonding between the atoms and hydrogen bonding between the molecules give water

the properties that make life possible. Presently, we are aware of no other planet that has liquid water.

In the next chapter, we will learn that a carbon atom combines covalently with CHNOPS to form the organic molecules of cells. It is these unique molecules that set living forms apart from nonliving objects. Carbon-containing molecules can be modified in numerous ways, and this accounts for life's diversity, such as differences between a bottle-nosed dolphin and a black shoulder peacock. Varying molecular compositions in plants can also tell us, for example, why

some trees have leaves that change color in the fall.

It is difficult for us to visualize that a bottle-nosed dolphin, a kangaroo, or a pine tree is a combination of molecules and ions, but later in this text we will learn that even our thoughts about these organisms are simply the result of molecules flowing from one brain cell to another. An atomic, ionic, and molecular understanding of the variety of processes unique to life provides a deeper understanding of the definition of life and offers tools for the improvement of its quality, preservation of its diversity, and appreciation of its beauty.

### summary

#### 2.1 Chemical Elements

Both living and nonliving things are composed of matter consisting of elements. The acronym CHNOPS stands for the most significant elements (atoms) found in living things: carbon, hydrogen, nitrogen, oxygen, phosphorus, and sulfur. Elements contain atoms, and atoms contain subatomic particles. Protons and neutrons in the nucleus determine the mass number of an atom. The atomic number indicates the number of protons and the number of electrons in electrically neutral atoms. Protons have positive charges, neutrons are uncharged, and electrons have negative charges. Isotopes are atoms of a single element that differ in their numbers of neutrons. Radioactive isotopes have many uses, including serving as tracers in biological experiments and medical procedures.

Electrons occupy energy levels (electron shells) at discrete distances from the nucleus. The number of electrons in the outer shell determines the reactivity of an atom. The first shell is complete when it is occupied by two electrons. In atoms up through calcium, number 20, every shell beyond the first shell is complete with eight electrons. The octet rule states that atoms react with one another in order to have a completed outer shell. Most atoms, including those common to living things, do not have filled outer shells and this causes them to react with one another to form compounds and/or molecules. Following the reaction, the atoms have completed outer shells.

#### 2.2 Compounds and Molecules

Ions form when atoms lose or gain one or more electrons to achieve a completed outer shell. An ionic bond is an attraction between oppositely charged ions. When covalent compounds form, atoms share electrons. A covalent bond is one or more shared pairs of electrons. There are single, double, and triple covalent bonds.

In polar covalent bonds, the sharing of electrons is not equal. If the molecule is polar, the more electronegative atom carries a slightly negative charge and the other atom carries a slightly positive charge.

#### 2.3 Chemistry of Water

Water is a polar molecule. The polarity of water molecules allows hydrogen bonding to occur between water molecules. A hydrogen

bond is a weak attraction between a slightly positive hydrogen atom and a slightly negative oxygen or nitrogen atom within the same or a different molecule. Hydrogen bonds help maintain the structure and function of cellular molecules.

Water's polarity and hydrogen bonding account for its unique properties. These features allow living things to exist and carry on cellular activities.

#### 2.4 Acids and Bases

A small fraction of water molecules dissociate to produce an equal number of hydrogen ions and hydroxide ions. Solutions with equal numbers of  $H^+$  and  $OH^-$  are termed neutral. In acidic solutions, there are more hydrogen ions than hydroxide ions; these solutions have a pH less than 7. In basic solutions, there are more hydroxide ions than hydrogen ions; these solutions have a pH greater than 7. Cells are sensitive to pH changes. Biological systems often contain buffers that help keep the pH within a normal range.

### understanding the terms

acid 32	hydroxide ion ( $OH^-$ ) 32
atom 22	ion 26
atomic mass 23	ionic bond 26
atomic number 23	isotope 24
atomic symbol 22	mass number 23
base 32	matter 22
buffer 32	molecule 26
calorie 29	neutron 22
compound 26	nonpolar covalent bond 27
covalent bond 27	octet rule 25
electron 22	pH scale 32
electronegativity 27	polar covalent bond 27
electron shell 22	proton 22
element 22	salt 26
evaporate 29	solute 29
formula 26	solution 29
hydrogen bond 28	surface tension 30
hydrogen ion ( $H^+$ ) 32	tracer 24
hydrophilic 30	valence shell 25
hydrophobic 30	

Match the terms to these definitions:

- \_\_\_\_\_ Bond in which the sharing of electrons between atoms is unequal.
- \_\_\_\_\_ Charged particle that carries a negative or positive charge(s).
- \_\_\_\_\_ Molecules tending to raise the hydrogen ion concentration in a solution and to lower its pH numerically.
- \_\_\_\_\_ The smallest part of a compound that still has the properties of that compound.
- \_\_\_\_\_ A chemical or a combination of chemicals that maintains a constant pH upon the addition of small amounts of acid or base.

## reviewing this chapter

1. Name the kinds of subatomic particles studied. What is their atomic mass unit, charge, and location in an atom? 21–23
2. What is an isotope? A radioactive isotope? Radioactivity is always considered dangerous. Why? 24–26
3. Using the Bohr model, draw an atomic structure for a carbon that has six protons and six neutrons. 26
4. Draw an atomic representation for  $\text{MgCl}_2$ . Using the octet rule, explain the structure of the compound. 27
5. Explain whether  $\text{CO}_2$  ( $\text{O}=\text{C}=\text{O}$ ) is an ionic or a covalent compound. Why does this arrangement satisfy all atoms involved? 27
6. Of what significance is the shape of molecules in organisms? 28
7. Explain why water is a polar molecule. What does the polarity and shape of water have to do with its ability to form hydrogen bonds? 28
8. Name five properties of water, and relate them to the structure of water, including its polarity and hydrogen bonding between molecules. 28–31
9. On the pH scale, which numbers indicate a solution is acidic? Basic? Neutral? 32
10. What are buffers, and why are they important to life? 32–33

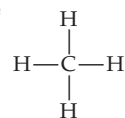
## testing yourself

Choose the best answer for each question.

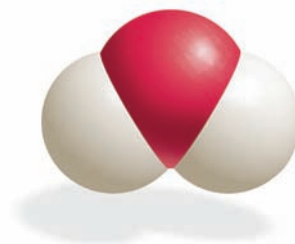
1. Which of the subatomic particles contributes almost no weight to an atom?
  - a. protons in the electron shells
  - b. electrons in the nucleus
  - c. neutrons in the nucleus
  - d. electrons at various energy levels
2. The atomic number tells you the
  - a. number of neutrons in the nucleus.
  - b. number of protons in the atom.
  - c. atomic mass of the atom.
  - d. number of its electrons if the atom is neutral.
  - e. Both b and d are correct.
3. An atom that has two electrons in the outer shell, such as magnesium, would most likely
  - a. share to acquire a completed outer shell.
  - b. lose these two electrons and become a negatively charged ion.
  - c. lose these two electrons and become a positively charged ion.
  - d. bind with carbon by way of hydrogen bonds.
  - e. bind with another calcium atom to satisfy its energy needs.

4. Isotopes differ in their
  - a. number of protons.
  - b. atomic number.
  - c. number of neutrons.
  - d. number of electrons.
5. When an atom gains electrons, it
  - a. forms a negatively charged ion.
  - b. forms a positively charged ion.
  - c. forms covalent bonds.
  - d. forms ionic bonds.
  - e. gains atomic mass.
6. A covalent bond is indicated by
  - a. plus and minus charges attached to atoms.
  - b. dotted lines between hydrogen atoms.
  - c. concentric circles about a nucleus.
  - d. overlapping electron shells or a straight line between atomic symbols.
  - e. the touching of atomic nuclei.
7. The shape of a molecule
  - a. is dependent in part on the angle of bonds between its atoms.
  - b. influences its biological function.
  - c. is dependent on its electronegativity.
  - d. is dependent on its place in the periodic table.
  - e. Both a and b are correct.
8. In which of these are the electrons always shared unequally?
  - a. double covalent bond
  - b. triple covalent bond
  - c. hydrogen bond
  - d. polar covalent bond
  - e. ionic and covalent bonds

9. In the molecule



- a. all atoms have eight electrons in the outer shell.
  - b. all atoms are sharing electrons.
  - c. carbon could accept more hydrogen atoms.
  - d. the bonds point to the corners of a square.
  - e. All of these are correct.
10. Which of these properties of water cannot be attributed to hydrogen bonding between water molecules?
    - a. Water stabilizes temperature inside and outside the cell.
    - b. Water molecules are cohesive.
    - c. Water is a solvent for many molecules.
    - d. Ice floats on liquid water.
    - e. Both b and c are correct.
  11. Complete this diagram by placing an O for oxygen or an H for hydrogen on the appropriate atoms. Place partial charges where they belong.





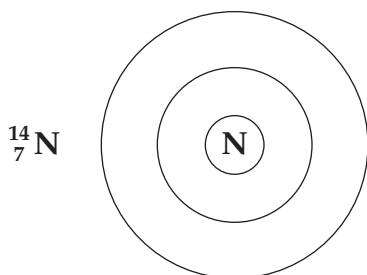
12.  $\text{H}_2\text{CO}_3/\text{NaHCO}_3$  is a buffer system in the body. What effect will the addition of an acid have on the pH of a solution that is buffered?
- The pH will rise.
  - The pH will lower.
  - The pH will not change.
  - All of these are correct.

13. Rainwater has a pH of about 5.6; therefore, rainwater is
- a neutral solution.
  - an acidic solution.
  - a basic solution.
  - It depends if it is buffered.

14. Acids
- release hydrogen ions in solution.
  - cause the pH of a solution to rise above 7.
  - take up hydroxide ions and become neutral.
  - increase the number of water molecules.
  - Both a and b are correct.

15. Which type of bond results from the sharing of electrons between atoms?
- covalent
  - ionic
  - hydrogen
  - neutral

16. Complete this diagram of a nitrogen atom by placing the correct number of protons and neutrons in the nucleus and electrons in the shells. Explain why the correct formula for ammonia is  $\text{NH}_3$ , not  $\text{NH}_4$ .



17. Why is  $\text{—NH}_2$  a polar group?
- Nitrogen is more electronegative than hydrogen.
  - The bonds are not symmetrical.
  - Because hydrogen bonding takes place.
  - Both a and b are correct.
18. If a chemical accepted  $\text{H}^+$  from the surrounding solution, the chemical could be
- a base.
  - an acid.
  - a buffer.
  - None of the above are correct.
  - Both a and c are correct.
19. The periodic table tells us
- the atomic number, symbol, and mass.
  - how many shells an atom has.
  - how many electrons are in the outer shell.
  - whether the atom will react or not.
  - All of these are correct.
20. Which of these best describes the changes that occur when a solution goes from pH 5 to pH 7?
- The solution is now 100 times more acidic.
  - The solution is now 100 times more basic.

- The hydrogen ion concentration decreases by only a factor of 20, as the solution goes from basic to acidic.
- The hydrogen ion concentration changes by only a factor of 20, as the solution goes from acidic to basic.

21. A hydrogen bond is not
- involved in maintaining the shape of certain molecules.
  - necessary to the properties of water.
  - as strong as a covalent bond.
  - represented by a dotted line.
  - More than one of these is correct.

For questions 22–25, match the statements with a property of water in the key.

#### KEY:

- Water flows because it is cohesive.
  - Water holds its heat.
  - Water has neutral pH.
  - Water has a high heat of vaporization.
22. Sweating helps cool us off.
23. Our blood is composed mostly of water and cells.
24. Our blood is just about pH 7.
25. We usually maintain a normal body temperature.

## thinking scientifically

- Natural phenomena often require an explanation. Based on how sodium chloride dissociates in water (see pages 29–30) and Figure 2.12, explain why the oceans don't freeze.
- Melvin Calvin used radioactive carbon (as a tracer) to discover a series of molecules that form during photosynthesis. Explain why carbon behaves chemically the same, even when radioactive.

## bioethical issue

### The Right to Refuse an IV

When a person gets sick or endures physical stress—as, for example, during childbirth—pH levels may dip or rise too far, endangering that person's life. In most U.S. hospitals, doctors routinely administer IVs, or intravenous infusions, of certain fluids to maintain a patient's pH level. Some people who oppose IVs for philosophical reasons may refuse an IV. That's relatively safe, as long as the person is healthy.

Problems arise when hospital policy dictates an IV, even though a patient does not want one. Should a patient be allowed to refuse an IV? Or does a hospital have the right to insist, for health reasons, that patients accept IV fluids? And what role should doctors play—patient advocates or hospital representatives?

## Biology website

The companion website for *Biology* provides a wealth of information organized and integrated by chapter. You will find practice tests, animations, videos, and much more that will complement your learning and understanding of general biology.

<http://www.mhhe.com/maderbiology10>

# 3

## The Chemistry of Organic Molecules

**W**e might have trouble thinking of ways that plants and animals are similar, but we all know that vegetarians have no trouble sustaining themselves by eating plants, as long as they include a variety of plants in their diet. That's because plants and humans generally have the same molecules in their cells—namely, carbohydrates, lipids, proteins, and nucleic acids. When we feed on plants, we digest their macromolecules to smaller molecules, and then we use these smaller molecules to build our own types of carbohydrates, lipids, proteins, and nucleic acids.

A similarity in chemistry between plants and humans is especially evident when we acquire vitamins from plants and use them exactly as plants do, because vitamins assist the same enzymes found in all organisms. The differences between plants and humans are due to their genes. But, then, all genes are made of DNA, and the way genes function in cells is the same in all organisms. In this chapter, we continue our look at basic chemistry by considering the molecules found in all living things. These are the types of molecules that account for the structure and function of all cells in any type of organism.

Granddaughter Sylvia with azalea plant.



### 3.1 ORGANIC MOLECULES

- All organic molecules have a skeleton composed of carbon chains. Variations in the carbon skeleton (straight, a ring, branched, or nonbranched) and the attached functional groups account for the great diversity of organic molecules. 38–39
- Large organic molecules form when their specific monomers join together. 40

### 3.2 CARBOHYDRATES

- Glucose, an immediate energy source, is the monomer for starch and glycogen, which are short-term stored energy sources. 41–43
- Other carbohydrates (cellulose and chitin) function as structural components of cells. 43

### 3.3 LIPIDS

- Lipids, nonsoluble in water, exist as fats (long-term energy storage), phospholipids (component of plasma membrane), steroids (e.g., hormones), and waxes (waterproof coverings). 44–47

### 3.4 PROTEINS

- Proteins, which have many and varied functions, are polymers of amino acids. Differences in levels of organization result in each protein having a particular shape. 48–51

### 3.5 NUCLEIC ACIDS

- Genes are composed of DNA (deoxyribonucleic acid). DNA specifies the correct ordering of amino acids in proteins, with RNA serving as an intermediary. 52–53
- The nucleotide ATP serves as a carrier of chemical energy in cells. 53–54



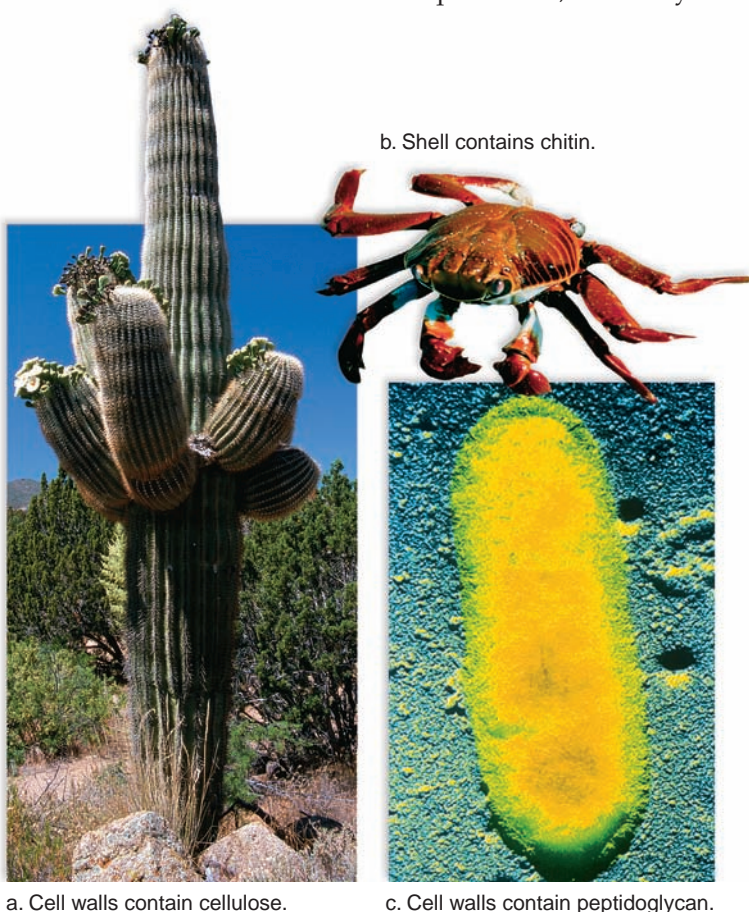
## 3.1 Organic Molecules

Because chemists of the nineteenth century thought that the molecules of cells must contain a vital force, they divided chemistry into **organic chemistry**, the chemistry of organisms, and **inorganic chemistry**, the chemistry of the nonliving world. This terminology is still with us even though many types of organic molecules can now be synthesized in the laboratory. Today, we simply define **organic molecules** as molecules that contain both carbon and hydrogen atoms (Table 3.1).

There are only four classes of organic compounds in any living thing: carbohydrates, lipids, proteins, and nucleic acids. Despite the limited number of classes, the so-called **biomolecules** in cells are quite diverse. A bacterial cell contains some 5,000 different organic molecules, and a plant or animal cell has twice that number. This diversity of organic molecules makes the diversity of life possible (Fig. 3.1). It is quite remarkable that the variety of organic molecules can be traced to the unique chemical properties of the carbon atom.

### The Carbon Atom

What is there about carbon that makes organic molecules the same and also different? Carbon is quite small, with only a



**FIGURE 3.1** Carbohydrates as structural materials.

**a.** Plants, such as cacti, have the carbohydrate cellulose in their cell walls. **b.** The shell of a crab contains chitin, a different carbohydrate. **c.** The cell walls of bacteria contain another type of carbohydrate known as peptidoglycan.

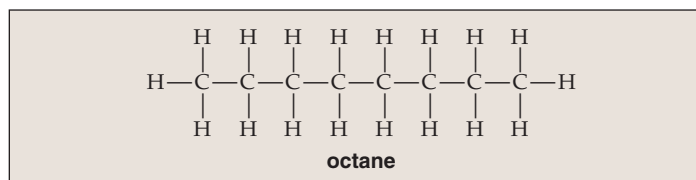
**TABLE 3.1**

**Inorganic Versus Organic Molecules**

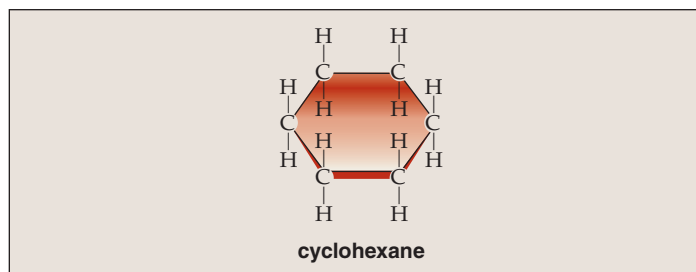
<i>Inorganic Molecules</i>	<i>Organic Molecules</i>
Usually contain positive and negative ions	Always contain carbon and hydrogen
Usually ionic bonding	Always covalent bonding
Always contain a small number of atoms	Often quite large, with many atoms
Often associated with nonliving matter	Usually associated with living organisms

total of six electrons: two electrons in the first shell and four electrons in the outer shell. In order to acquire four electrons to complete its outer shell, a carbon atom almost always shares electrons with—you guessed it—CHNOPS, the elements that make up most of the weight of living things (see Fig. 2.1).

Because carbon needs four electrons to complete its outer shell, it can share with as many as four other elements, and this spells diversity. But even more significant to the shape, and therefore the function, of biomolecules, carbon often shares electrons with another carbon atom. The C—C bond is quite stable, and the result is carbon chains that can be quite long. Hydrocarbons are chains of carbon atoms bonded exclusively to hydrogen atoms.



Branching at any carbon atom is possible, and also a hydrocarbon can turn back on itself to form a ring compound when placed in water:



Carbon can form double bonds with itself and other atoms. Double bonds restrict the movement of bonded atoms, and in that way contribute to the shape of the molecule. As in acetylene,  $\text{H}-\text{C}\equiv\text{C}-\text{H}$ , carbon is also capable of forming a triple bond with itself.

The diversity of organic molecules is further enhanced by the presence of particular functional groups, as discussed next. Contrast the structure of cyclohexane, above, with the structure of glucose in Figure 3.6. The difference in structure can be attributed to the functional groups added to the same number of carbons.

## The Carbon Skeleton and Functional Groups

The carbon chain of an organic molecule is called its skeleton or backbone. The terminology is appropriate because just as a skeleton accounts for your shape, so does the carbon skeleton of an organic molecule account for its shape. Vertebrates look very different, even though they all have a backbone of vertebrae. We recognize them by their shape and also by their appendages, whether they have fins, wings, or limbs, for example. So, the diversity of organic molecules comes about when different functional groups are added to the carbon skeleton. A **functional group** is a specific combination of bonded atoms that always re-

acts in the same way, regardless of the particular carbon skeleton. As in Figure 3.2, it is even acceptable to use an *R* to stand for the remainder of the molecule, which is the carbon skeleton, because only the functional group is involved in a reaction.

Notice that when a particular functional group is added to a carbon skeleton, the molecule becomes a certain type of compound. For example, the addition of an  $\text{—OH}$  (hydroxyl group) to a carbon skeleton turns that molecule into an alcohol. When an  $\text{—OH}$  replaces one of the hydrogens in ethane, a 2-carbon hydrocarbon, it becomes ethanol, a type of alcohol that is familiar because it is consumable by humans. Whereas ethane, like other hydrocarbons, is **hydrophobic** (not soluble in water), ethanol is **hydrophilic** (soluble in water) because the  $\text{—OH}$  functional group is polar. Since cells are 70–90% water, the ability to interact with and be soluble in water profoundly affects the function of organic molecules in cells.

Organic molecules containing carboxyl (acidic) groups ( $\text{—COOH}$ ) are highly polar. They tend to ionize and release hydrogen ions in solution:



The attached functional groups determine the polarity of an organic molecule and also the types of reactions it will undergo. We will see that alcohols react with carboxyl groups when a fat forms, and that carboxyl groups react with amino groups during protein formation.

### Isomers

**Isomers** [Gk. *isos*, equal, and *meros*, part, portion] are organic molecules that have identical molecular formulas but a different arrangement of atoms. In essence, isomers are variations in the molecular architecture of a molecule. Isomers are another example of how the chemistry of carbon leads to variations in organic molecules.

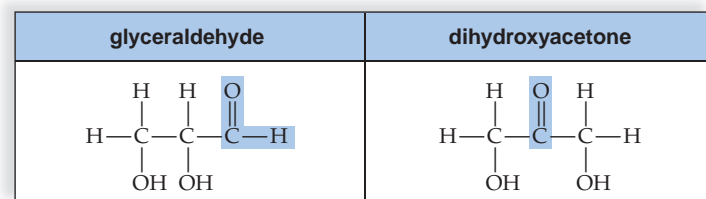
The two molecules in Figure 3.3 are isomers of one another; they have the same molecular formula but different functional groups. Therefore, we would expect them to react differently in chemical reactions.

Functional Groups			
Group	Structure	Compound	Significance
Hydroxyl	$\text{R—OH}$	Alcohol as in ethanol	Polar, forms hydrogen bond, present in sugars and some amino acids
Carbonyl	$\text{R—}\overset{\text{O}}{\parallel}\text{C—H}$	Aldehyde as in formaldehyde	Polar, present in sugars
	$\text{R—}\overset{\text{O}}{\parallel}\text{C—R}$	Ketone as in acetone	
Carboxyl (acidic)	$\text{R—}\overset{\text{O}}{\parallel}\text{C—OH}$	Carboxylic acid as in acetic acid	Polar, acidic, present in fatty acids and amino acids
Amino	$\text{R—N}\begin{matrix} \text{H} \\ \diagup \diagdown \\ \text{H} \end{matrix}$	Amine as in tryptophan	Polar, basic, forms hydrogen bonds, present in amino acids
Sulfhydryl	$\text{R—SH}$	Thiol as in ethanethiol	Forms disulfide bonds, present in some amino acids
Phosphate	$\text{R—O—}\overset{\text{O}}{\parallel}\text{P—OH}$ OH	Organic phosphate as in phosphorylated molecules	Polar, acidic, present in nucleotides and phospholipids

R = remainder of molecule

**FIGURE 3.2** Functional groups.

Molecules with the same carbon skeleton can still differ according to the type of functional group attached to the carbon skeleton. Many of these functional groups are polar, helping to make the molecule soluble in water. In this illustration, the remainder of the molecule (does not include the functional group) is represented by an *R*.



**FIGURE 3.3** Isomers.

Isomers have the same molecular formula but different atomic configurations. Both of these compounds have the formula  $\text{C}_3\text{H}_6\text{O}_3$ . In glyceraldehyde, oxygen is double-bonded to an end carbon. In dihydroxyacetone, oxygen is double-bonded to the middle carbon.





Biomolecules		
Category	Example	Subunit(s)
Carbohydrates*	Polysaccharide	Monosaccharide
Lipids	Fat	Glycerol and fatty acids
Proteins*	Polypeptide	Amino acids
Nucleic acids*	DNA, RNA	Nucleotide

\*Polymers

**FIGURE 3.4** Common foods.

Carbohydrates in bread and pasta are digested to sugars; lipids such as oils are digested to glycerol and fatty acids; and proteins in meat are digested to amino acids. Cells use these subunit molecules to build their own biomolecules and as a source of energy.

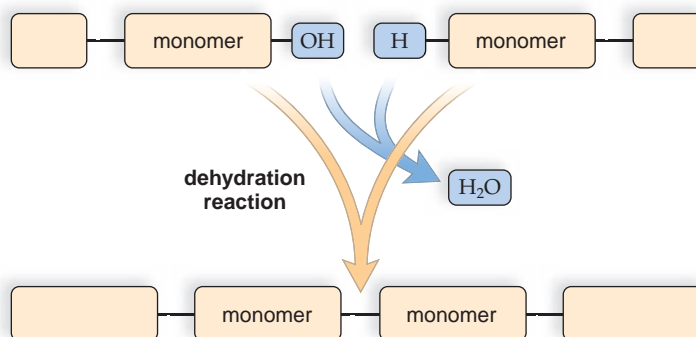
## The Biomolecules of Cells

You are very familiar with the names of biomolecules—carbohydrates, lipids, proteins, and nucleic acids—because certain foods are known to be rich in them, as illustrated in Figure 3.4. For example, bread is rich in carbohydrate, and meat is rich in protein. When you digest food, it gets broken down into smaller molecules that are subunits for biomolecules. Digestion of bread releases glucose molecules, digestion of meat releases amino acids. Your body then takes these subunits and builds from them the particular carbohydrates and proteins that make up your cells (Fig. 3.4, *below*).

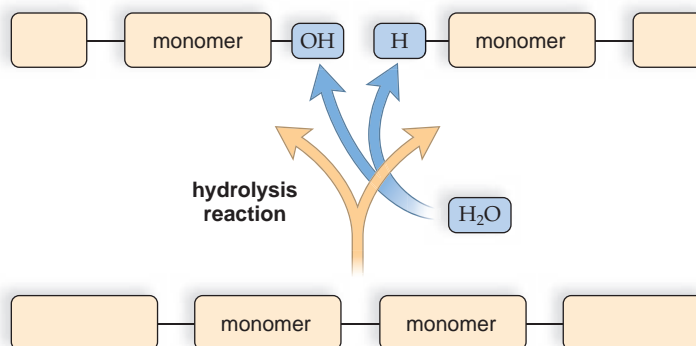
### Synthesis and Degradation

A cell uses a condensation reaction to synthesize (build up) any type of biomolecule. It's called a **dehydration reaction** because the equivalent of a water molecule—that is, an —OH (hydroxyl group) and an —H (hydrogen atom), is removed as subunits are joined. Therefore, water molecules result as biomolecules are synthesized (Fig. 3.5a).

To break down biomolecules, a cell uses an opposite type of reaction. During a **hydrolysis** [Gk. *hydro*, water, and *lyse*, break] **reaction**, an —OH group from water attaches to one subunit, and an —H from water attaches to the other subunit. In other words, biomolecules are broken down by adding water to them (Fig. 3.5b).



a. Synthesis of a biomolecule



b. Degradation of a biomolecule

**FIGURE 3.5** Synthesis and degradation of biomolecules.

**a.** In cells, synthesis often occurs when subunits bond during a dehydration reaction (removal of  $H_2O$ ). **b.** Degradation occurs when the subunits separate during a hydrolysis reaction (the addition of  $H_2O$ ).

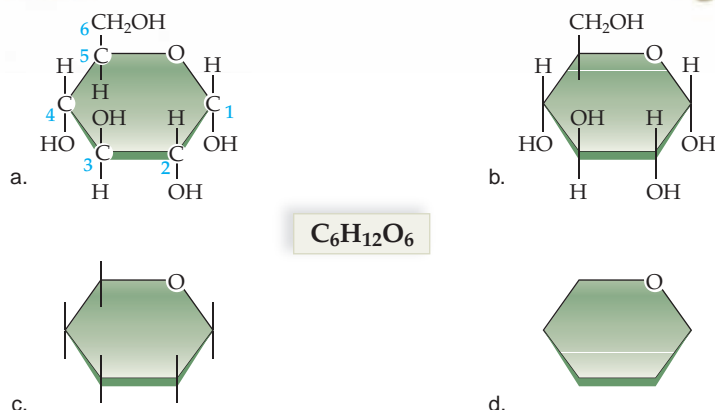
Enzymes are required for cells to carry out dehydration and hydrolysis reactions. An **enzyme** is a molecule that speeds a reaction by bringing reactants together, and the enzyme may even participate in the reaction but it is unchanged by it.

**Polymers.** The largest of the biomolecules are called **polymers** and like all biomolecules, polymers are constructed by linking together a large number of the same type of subunit. However, in the case of polymers, the subunits are called **monomers**. A polysaccharide, a protein and a nucleic acid, is a polymer that contains innumerable monomers. Just as a train increases in length when boxcars are hitched together one by one, so a polymer gets longer as monomers bond to one another.

### Check Your Progress

3.1

- Describe the properties of a carbon atom that make it ideally suited to produce varied carbon skeletons.
- a.** How could two pearl necklaces be both the same and different? **b.** How could two protein polymers be both the same and different?



**FIGURE 3.6** Glucose.

Glucose provides energy for organisms, such as this cheetah. Each of these structural formulas is glucose. **a.** The carbon skeleton and all attached groups are shown. **b.** The carbon skeleton is omitted. **c.** The carbon skeleton and attached groups are omitted. **d.** Only the ring shape, which includes one oxygen atom, remains.

## 3.2 Carbohydrates

**Carbohydrates** are almost universally used as an immediate energy source in living things, but they also play structural roles in a variety of organisms (see Fig. 3.1). The majority of carbohydrates have a carbon to hydrogen to oxygen ratio of 1:2:1. The term *carbohydrate* includes single sugar molecules and also chains of sugars. Chain length varies from a few sugars to hundreds of sugars. The long chains are thus polymers. The monomers of carbohydrates are monosaccharides.

### Monosaccharides: Ready Energy

**Monosaccharides** [Gk. *monos*, single, and *sacchar*, sugar], consisting of only a single sugar molecule, are called simple sugars. A simple sugar can have a carbon backbone of three to

seven carbons. The molecular formula for a simple sugar is some multiple of  $CH_2O$ , suggesting that every carbon atom is bonded to an  $-H$  and an  $-OH$ . This is not strictly correct, as you can see by examining the structural formula for glucose (Fig. 3.6). Still, sugars do have many hydroxyl groups, and this polar functional group makes them soluble in water.

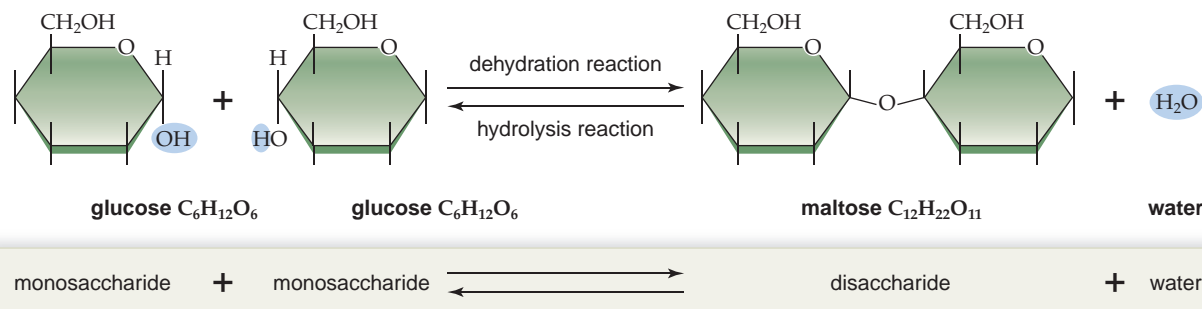
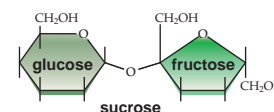
**Glucose**, with six carbon atoms, is a **hexose** [Gk. *hex*, six] and has a molecular formula of  $C_6H_{12}O_6$ . Despite the fact that glucose has several isomers, such as fructose and galactose, we usually think of  $C_6H_{12}O_6$  as glucose. This signifies that glucose has a special place in the chemistry of organisms. This simple sugar is the major source of cellular fuel for all living things. Glucose is transported in the blood of animals, and it is the molecule that is broken down in nearly all types of organisms during cellular respiration, with the resulting buildup of ATP molecules.

**Ribose** and **deoxyribose**, with five carbon atoms, are **pentoses** [Gk. *pent*, five] of significance because they are found respectively in the nucleic acids RNA and DNA. RNA and DNA are discussed later in the chapter.

### Disaccharides: Varied Uses

A **disaccharide** contains two monosaccharides that have joined during a dehydration reaction. Figure 3.7 shows how the disaccharide maltose (an ingredient used in brewing) arises when two glucose molecules bond together. Note the position of the bond that results when the  $-OH$  groups participating in the reaction project below the ring. When our hydrolytic digestive juices break this bond, the result is two glucose molecules.

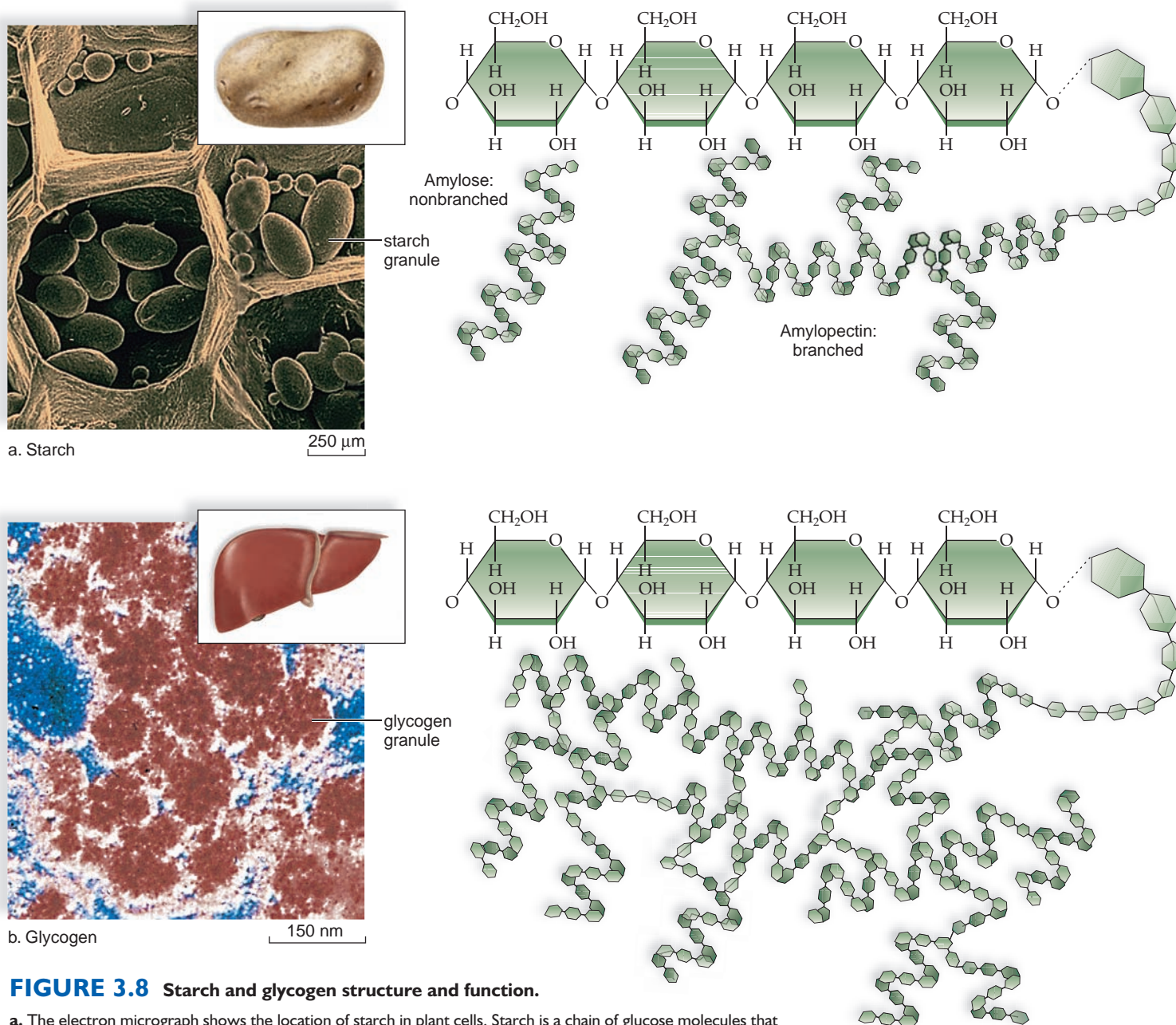
Sucrose (the structure shown at right) is another disaccharide of special interest because it is sugar we use at home to sweeten our food. Sucrose is also the form in which sugar is transported in plants. We acquire sucrose from plants such as sugarcane and sugar beets. You may also have heard of lactose, a disaccharide found in milk. Lactose is glucose combined with galactose. Individuals that are lactose intolerant cannot break this disaccharide down and have subsequent medical problems. To prevent problems they can buy foods in which lactose has been broken down into its subunits.



**FIGURE 3.7** Synthesis and degradation of maltose, a disaccharide.

Synthesis of maltose occurs following a dehydration reaction when a bond forms between two glucose molecules, and water is removed. Degradation of maltose occurs following a hydrolysis reaction when this bond is broken by the addition of water.





## Polysaccharides: Energy Storage Molecules

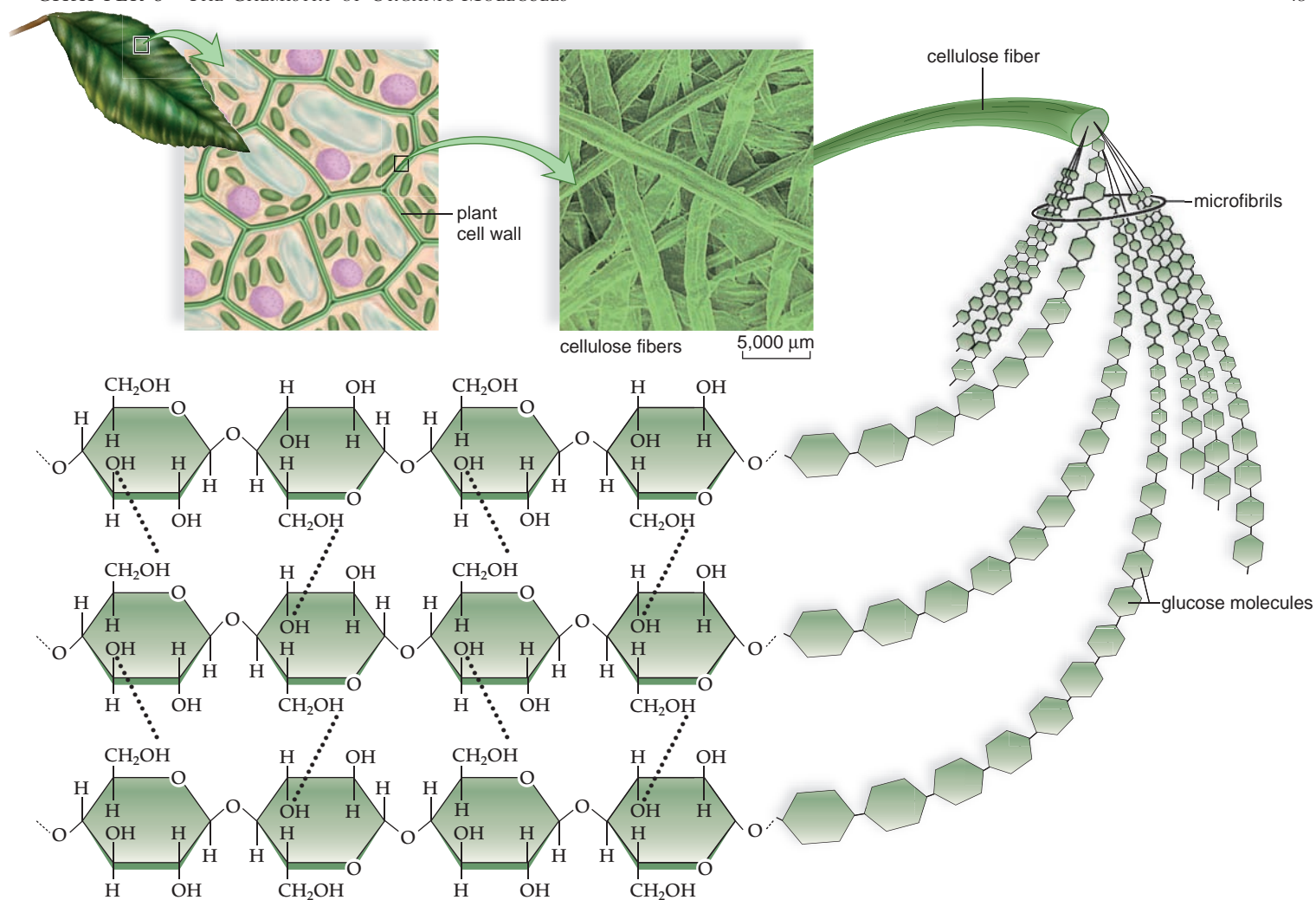
**Polysaccharides** are polymers of monosaccharides. Some types of polysaccharides function as short-term energy storage molecules. When an organism requires energy, the polysaccharide is broken down to release sugar molecules. The helical shape of the polysaccharides in Figure 3.8 exposes the sugar linkages to the hydrolytic enzymes that can break them down.

Plants store glucose as **starch**. The cells of a potato contain granules where starch resides during winter until energy is needed for growth in the spring. Notice in Figure 3.8a that starch exists in two forms: One form (amylose) is nonbranched and the other (amylopectin) is branched.

When a polysaccharide is branched, there is no main carbon chain because new chains occur at regular intervals, always at the sixth carbon of the monomer.

Animals store glucose as **glycogen**. In our bodies and those of other vertebrates, liver cells contain granules where glycogen is being stored until needed. The storage and release of glucose from liver cells is under the control of hormones. After we eat, the release of the hormone insulin from the pancreas promotes the storage of glucose as glycogen. Notice in Figure 3.8b that glycogen is even more branched than starch.

Polysaccharides serve as storage molecules because they are not as soluble in water, and are much larger than a sugar. Therefore, polysaccharides cannot easily pass through the plasma membrane, a sheetlike structure that encloses cells.



**FIGURE 3.9** Cellulose fibrils.

Cellulose fibers criss-cross in plant cell walls for added strength. A cellulose fiber contains several microfibrils, each a polymer of glucose molecules—notice that the linkage bonds differ from those of starch. Every other glucose is flipped, permitting hydrogen bonding between the microfibrils.

## Polysaccharides: Structural Molecules

Structural polysaccharides include **cellulose** in plants, **chitin** in animals and fungi, and **peptidoglycan** in bacteria (see Fig. 3.1). In all three, monomers are joined by the type of bond shown for cellulose in Figure 3.9. The cellulose monomer is simply glucose, but in chitin, the monomer has an attached amino group. The structure of peptidoglycan is even more complex because each monomer also has an amino acid chain.

Cellulose is the most abundant carbohydrate and, indeed, the most abundant organic molecule on Earth—over 100 billion tons of cellulose is produced by plants each year. Wood, a cellulose plant product, is used for construction, and cotton is used for cloth. Microorganisms, but not animals, are able to digest the bond between glucose monomers in cellulose. The protozoans in the gut of termites allows termites to digest wood. In cows and other ruminants, microorganisms break down cellulose in a special pouch before the “cud” is returned to the mouth for more chewing and reswallowing. In rabbits, microorganisms digest cellulose in a pouch where it is packaged into pellets. In order to make use of these nutrient pellets, rabbits have to reswallow them

as soon as they pass out at the anus. For animals, such as humans, that have no means of digesting cellulose, cellulose is dietary fiber, which maintains regularity of elimination.

Chitin [Gk. *chiton*, tunic] is found in fungal cell walls and in the exoskeletons of crabs and related animals, such as lobsters, scorpions, and insects. Chitin, like cellulose, cannot be digested by animals; however, humans have found many other good uses for chitin. Seeds are coated with chitin, and this protects them from attack by soil fungi. Because chitin also has antibacterial and antiviral properties, it is processed and used in medicine as a wound dressing and suture material. Chitin is even useful during the production of cosmetics and various foods.

## Check Your Progress

### 3.2

1. Explain why humans cannot utilize the glucose in cellulose as a nutrient source.
2. Compare and contrast the structure and function of cellulose with chitin.



### 3.3 Lipids

A variety of organic compounds are classified as **lipids** [Gk. *lipos*, fat] (Table 3.2). These compounds are insoluble in water due to their hydrocarbon chains. Hydrogens bonded only to carbon have no tendency to form hydrogen bonds with water molecules. Fat, a well-known lipid, is used for both insulation and long-term energy storage by animals. Fat below the skin of marine mammals is called blubber (Fig. 3.10); in humans, it is given slang expressions such as “spare tire” and “love handles.” Plants use oil instead of fat for long-term energy storage. We are familiar with fats and oils because we use them as foods and for cooking.

Phospholipids and steroids are also important lipids found in living things. They serve as major components of the plasma membrane in cells. Waxes, which are sticky, not greasy like fats and oils, tend to have a protective function in living things.

#### Triglycerides: Long-Term Energy Storage

**Fats** and **oils** contain two types of subunit molecules: fatty acids and glycerol. Each **fatty acid** consists of a long hydrocarbon chain with a  $\text{—COOH}$  (carboxyl) group at one end. Most of the fatty acids in cells contain 16 or 18 carbon atoms per molecule, although smaller ones are also found. Fatty acids are either saturated or unsaturated. **Saturated fatty acids** have no double bonds between the carbon atoms. The carbon chain is saturated, so to speak, with all the hydrogens that can be held. **Unsaturated fatty acids** have double bonds in the carbon chain wherever the number of hydrogens is less than two per carbon atom.

**Glycerol** is a compound with three  $\text{—OH}$  groups. The  $\text{—OH}$  groups are polar; therefore, glycerol is soluble in water. When a fat or oil forms, the acid portions of three fatty acids react with the  $\text{—OH}$  groups of glycerol during a de-

**FIGURE 3.10 Blubber.**

The fat (blubber) beneath the skin of marine mammals protects them well from the cold. Blubber accounts for about 25% of their body weight.



hydration reaction (Fig. 3.11a). In addition to a fat molecule, three molecules of water result. Fats and oils are degraded following a hydrolysis reaction. Because there are three fatty acids attached to each glycerol molecule, fats and oils are sometimes called **triglycerides**. Notice that triglycerides have many  $\text{C—H}$  bonds; therefore, they do not mix with water. Despite the liquid nature of both cooking oils and water, cooking oils separate out of water even after shaking.

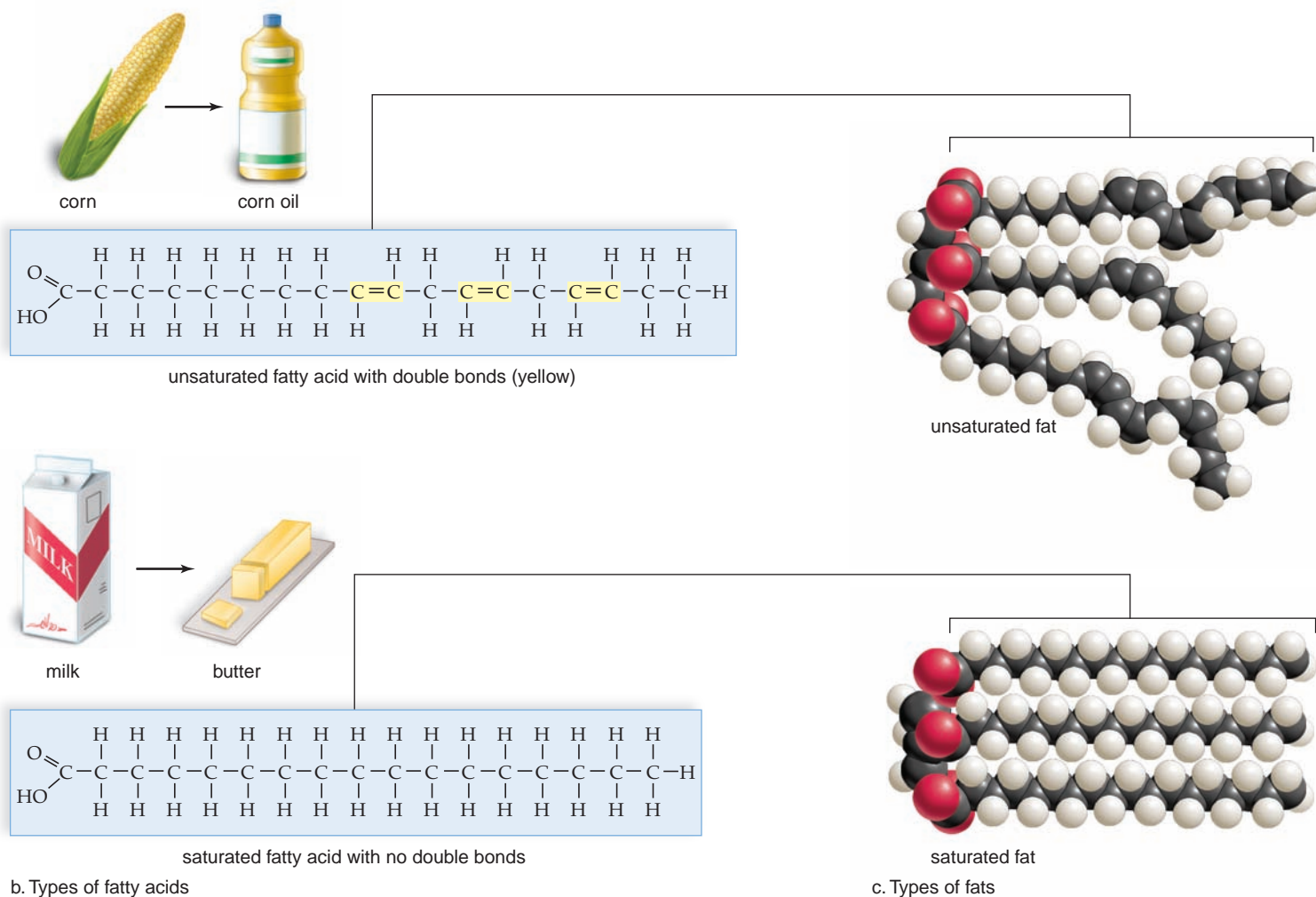
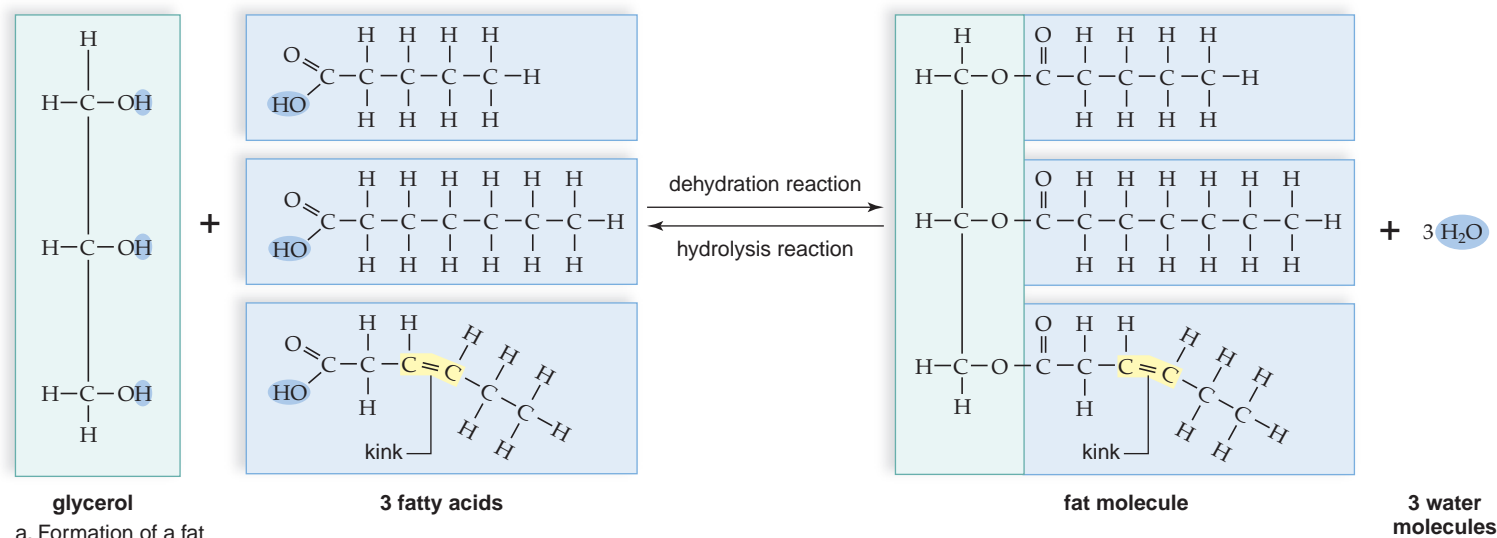
Triglycerides containing fatty acids with unsaturated bonds melt at a lower temperature than those containing only saturated fatty acids. This is because a double bond creates a kink in the fatty acid chain that prevents close packing between the hydrocarbon chains (Fig. 3.11a). We can reason that butter, a fat that is solid at room temperature, must contain primarily saturated fatty acids, while corn oil, which is a liquid even when placed in the refrigerator, must contain primarily unsaturated fatty acids (Fig. 3.11b). This difference is useful to living things. For example, the feet of reindeer and penguins contain unsaturated triglycerides, and this helps protect those exposed parts from freezing.

In general, however, fats, which are most often of animal origin, are solid at room temperature, and oils, which are liquid at room temperature, are of plant origin. Diets high in animal fat have been associated with circulatory disorders because fatty material accumulates inside the lining of blood vessels and blocks blood flow. Replacement of fat whenever possible with oils such as olive oil and canola oil has been suggested.

Nearly all animals use fat in preference to glycogen for long-term energy storage. Gram per gram, fat stores more energy than glycogen. The  $\text{C—H}$  bonds of fatty acids make them a richer source of chemical energy than glycogen, because glycogen has many  $\text{C—OH}$  bonds. Also, fat droplets, being nonpolar, do not contain water. Small birds, like the broad-tailed hummingbird, store a great deal of fat before they start their long spring and fall migratory flights. About 0.15 g of fat per gram of body weight is accumulated each day. If the same amount of energy were stored as glycogen, a bird would be so heavy it would not be able to fly.

**TABLE 3.2**

Lipids		
Type	Functions	Human Uses
Fats	Long-term energy storage and insulation in animals	Butter, lard
Oils	Long-term energy storage in plants and their seeds	Cooking oils
Phospholipids	Component of plasma membrane	—
Steroids	Component of plasma membrane (cholesterol), sex hormones	Medicines
Waxes	Protection, prevent water loss (cuticle of plant surfaces), beeswax, earwax	Candles, polishes



**FIGURE 3.11 Fat and fatty acids.**

**a.** Following a dehydration reaction, glycerol is bonded to three fatty acid molecules as fat forms and water is given off. Following a hydrolysis reaction, the bonds are broken due to the addition of water. **b.** A fatty acid has a carboxyl group attached to a long hydrocarbon chain. If there are double bonds between some of the carbons in the chain, the fatty acid is unsaturated and a kink occurs in the chain. If there are no double bonds, the fatty acid is saturated. **c.** Space-filling models of an unsaturated fat and a saturated fat.