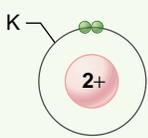
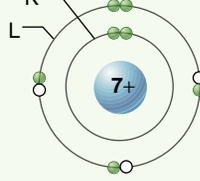


Nonreactive	Reactive
2 protons 2 neutrons 2 electrons	7 protons 7 neutrons 7 electrons
	
Helium	Nitrogen

**Figure 2.8** Electron energy levels for helium and nitrogen. Green balls represent electrons, blue ball represents the nucleus with number of protons indicated by number of (+) charges. Note that the helium atom has a filled K shell and is thus unreactive, whereas the nitrogen atom has five electrons in the L shell, three of which are unpaired, making it reactive.

These organic compounds contain primarily these four elements (CHON), explaining their prevalence in living systems. Some trace elements, such as zinc (Zn) and iodine (I), play crucial roles in living processes even though they are present in tiny amounts. Iodine deficiency, for example, can lead to enlargement of the thyroid gland, causing a bulge at the neck called a goiter.

### Learning Outcomes Review 2.2

The periodic table shows the elements in terms of atomic number and repeating chemical properties. Only 12 elements are found in significant amounts in living organisms: C, H, O, N, P, S, Na, K, Ca, Mg, Fe, and Cl.

- Why are the noble gases more stable than other elements in the periodic table?

## 2.3 The Nature of Chemical Bonds

### Learning Outcomes

1. Relate position in the periodic table to the formation of ions.
2. Explain how complex molecules can be built from many atoms by covalent bonds.
3. Contrast polar and nonpolar covalent bonds.

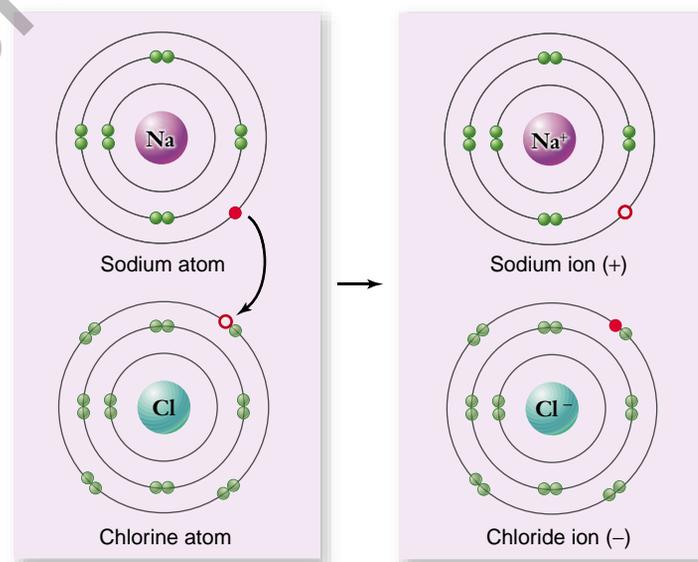
A group of atoms held together by energy in a stable association is called a *molecule*. When a molecule contains atoms of more than one element, it is called a *compound*. The atoms in a molecule are joined by *chemical bonds*; these bonds can result when atoms with opposite charges attract each other (ionic bonds), when two atoms share one or more pairs of electrons

TABLE 2.1 Bonds and Interactions		
Name	Basis of Interaction	Strength
Covalent bond	Sharing of electron pairs	Strong
Ionic bond	Attraction of opposite charges	
Hydrogen bond	Sharing of H atom	Weak
Hydrophobic interaction	Forcing of hydrophobic portions of molecules together in presence of polar substances	
van der Waals attraction	Weak attractions between atoms due to oppositely polarized electron clouds	Weak

(covalent bonds), or when atoms interact in other ways (table 2.1). We will start by examining *ionic bonds*, which form when atoms with opposite electrical charges (ions) attract.

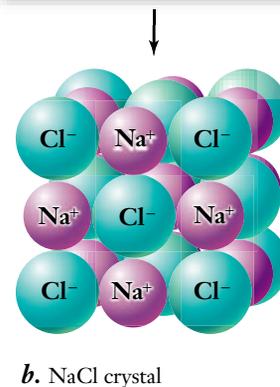
### Ionic bonds form crystals

Common table salt, the molecule sodium chloride (NaCl), is a lattice of ions in which the atoms are held together by ionic bonds (figure 2.9). Sodium has 11 electrons: 2 in the inner



**Figure 2.9** The formation of ionic bonds by sodium chloride.

*a.* When a sodium atom donates an electron to a chlorine atom, the sodium atom becomes a positively charged sodium ion, and the chlorine atom becomes a negatively charged chloride ion. *b.* The electrostatic attraction of oppositely charged ions leads to the formation of a lattice of  $\text{Na}^+$  and  $\text{Cl}^-$ .



energy level (K), 8 in the next level (L), and 1 in the outer (valence) level (M). The single, unpaired valence electron has a strong tendency to join with another unpaired electron in another atom. A stable configuration can be achieved if the valence electron is lost to another atom that also has an unpaired electron. The loss of this electron results in the formation of a positively charged sodium ion,  $\text{Na}^+$ .

The chlorine atom has 17 electrons: 2 in the K level, 8 in the L level, and 7 in the M level. As you can see in the figure, one of the orbitals in the outer energy level has an unpaired electron (red circle). The addition of another electron fills that level and causes a negatively charged chloride ion,  $\text{Cl}^-$ , to form.

When placed together, metallic sodium and gaseous chlorine react swiftly and explosively, as the sodium atoms donate electrons to chlorine to form  $\text{Na}^+$  and  $\text{Cl}^-$  ions. Because opposite charges attract, the  $\text{Na}^+$  and  $\text{Cl}^-$  remain associated in an *ionic compound*,  $\text{NaCl}$ , which is electrically neutral. The electrical attractive force holding  $\text{NaCl}$  together, however, is not directed specifically between individual  $\text{Na}^+$  and  $\text{Cl}^-$  ions, and no individual sodium chloride molecules form. Instead, the force exists between any one ion and *all* neighboring ions of the opposite charge. The ions aggregate in a crystal matrix with a precise geometry. Such aggregations are what we know as salt crystals. If a salt such as  $\text{NaCl}$  is placed in water, the electrical attraction of the water molecules, for reasons we will point out later in this chapter, disrupts the forces holding the ions in their crystal matrix, causing the salt to dissolve into a roughly equal mixture of free  $\text{Na}^+$  and  $\text{Cl}^-$  ions.

Because living systems always include water, ions are more important than ionic crystals. Important ions in biological systems include  $\text{Ca}^{2+}$ , which is involved in cell signaling,  $\text{K}^+$  and  $\text{Na}^+$ , which are involved in the conduction of nerve impulses.

## Covalent bonds build stable molecules

*Covalent bonds* form when two atoms share one or more pairs of valence electrons. Consider gaseous hydrogen ( $\text{H}_2$ ) as an example. Each hydrogen atom has an unpaired electron and an unfilled outer energy level; for these reasons, the hydrogen atom is unstable. However, when two hydrogen atoms are in close association, each atom's electron is attracted to both nuclei. In effect, the nuclei are able to share their electrons. The result is a diatomic (two-atom) molecule of hydrogen gas.

The molecule formed by the two hydrogen atoms is stable for three reasons:

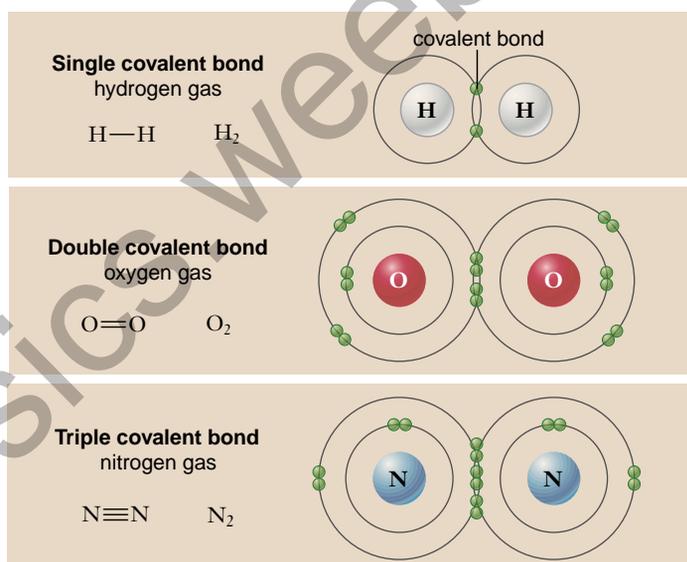
1. **It has no net charge.** The diatomic molecule formed as a result of this sharing of electrons is not charged because it still contains two protons and two electrons.
2. **The octet rule is satisfied.** Each of the two hydrogen atoms can be considered to have two orbiting electrons in its outer energy level. This state satisfies the octet rule, because each shared electron orbits both nuclei and is included in the outer energy level of both atoms.
3. **It has no unpaired electrons.** The bond between the two atoms also pairs the two free electrons.

Unlike ionic bonds, covalent bonds are formed between two individual atoms, giving rise to true, discrete molecules.

## The strength of covalent bonds

The strength of a covalent bond depends on the number of shared electrons. Thus *double bonds*, which satisfy the octet rule by allowing two atoms to share two pairs of electrons, are stronger than *single bonds*, in which only one electron pair is shared. In practical terms, more energy is required to break a double bond than a single bond. The strongest covalent bonds are *triple bonds*, such as those that link the two nitrogen atoms of nitrogen gas molecules ( $\text{N}_2$ ).

Covalent bonds are represented in chemical formulas as lines connecting atomic symbols. Each line between two bonded atoms represents the sharing of one pair of electrons. The *structural formulas* of hydrogen gas and oxygen gas are  $\text{H}-\text{H}$  and  $\text{O}=\text{O}$ , respectively, and their *molecular formulas* are  $\text{H}_2$  and  $\text{O}_2$ . The structural formula for  $\text{N}_2$  is  $\text{N}\equiv\text{N}$ .



## Molecules with several covalent bonds

A vast number of biological compounds are composed of more than two atoms. An atom that requires two, three, or four additional electrons to fill its outer energy level completely may acquire them by sharing its electrons with two or more other atoms.

For example, the carbon atom (C) contains six electrons, four of which are in its outer energy level and are unpaired. To satisfy the octet rule, a carbon atom must form four covalent bonds. Because four covalent bonds may form in many ways, carbon atoms are found in many different kinds of molecules.  $\text{CO}_2$  (carbon dioxide),  $\text{CH}_4$  (methane), and  $\text{C}_2\text{H}_5\text{OH}$  (ethanol) are just a few examples.

## Polar and nonpolar covalent bonds

Atoms differ in their affinity for electrons, a property called **electronegativity**. In general, electronegativity increases left to right across a row of the periodic table and decreases down the column. Thus the elements in the upper-right corner have the highest electronegativity.

For bonds between identical atoms, for example, between two hydrogen or two oxygen atoms, the affinity for electrons is obviously the same, and the electrons are equally shared. Such

bonds are termed **nonpolar**. The resulting compounds (H<sub>2</sub> or O<sub>2</sub>) are also referred to as nonpolar.

For atoms that differ greatly in electronegativity, electrons are not shared equally. The shared electrons are more likely to be closer to the atom with greater electronegativity, and less likely to be near the atom of lower electronegativity. In this case, although the molecule is still electrically neutral (same number of protons as electrons), the distribution of charge is not uniform. This unequal distribution results in regions of partial negative charge near the more electronegative atom, and regions of partial positive charge near the less electronegative atom. Such bonds are termed **polar covalent bonds**, and the molecules polar molecules. When drawing polar molecules, these partial charges are usually symbolized by the lowercase Greek letter delta ( $\delta$ ). The partial charge seen in a polar covalent bond is relatively small—far less than the unit charge of an ion. For biological molecules, we can predict polarity of bonds by knowing the relative electronegativity of a small number of important atoms (table 2.2). Notice that although C and H differ slightly in electronegativity, this small difference is negligible, and C–H bonds are considered nonpolar.

Because of its importance in the chemistry of water, we will explore the nature of polar and nonpolar molecules in the following section on water. Water (H<sub>2</sub>O) is a polar molecule with electrons more concentrated around the oxygen atom.

## Chemical reactions alter bonds

The formation and breaking of chemical bonds, which is the essence of chemistry, is termed a *chemical reaction*. All chemical reactions involve the shifting of atoms from one molecule or ionic compound to another, without any change in the number or identity of the atoms. For convenience, we refer to the original molecules before the reaction starts as *reactants*, and the molecules resulting from the chemical reaction as *products*. For example:



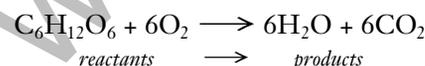
You may recognize this reaction as a simplified form of the photosynthesis reaction, in which water and carbon dioxide are combined to produce glucose and oxygen. Most animal life ultimately depends on this reaction, which takes place in plants. (Photosynthetic reactions will be discussed in detail in chapter 8.)

TABLE 2.2		Relative Electronegativities of Some Important Atoms
Atom		Electronegativity
O		3.5
N		3.0
C		2.5
H		2.1

The extent to which chemical reactions occur is influenced by three important factors:

- 1. Temperature.** Heating the reactants increases the rate of a reaction because the reactants collide with one another more often. (Care must be taken that the temperature is not so high that it destroys the molecules.)
- 2. Concentration of reactants and products.** Reactions proceed more quickly when more reactants are available, allowing more frequent collisions. An accumulation of products typically slows the reaction and, in reversible reactions, may speed the reaction in the reverse direction.
- 3. Catalysts.** A catalyst is a substance that increases the rate of a reaction. It doesn't alter the reaction's equilibrium between reactants and products, but it does shorten the time needed to reach equilibrium, often dramatically. In living systems, proteins called enzymes catalyze almost every chemical reaction.

Many reactions in nature are reversible. This means that the products may themselves be reactants, allowing the reaction to proceed in reverse. We can write the preceding reaction in the reverse order:



This reaction is a simplified version of the oxidation of glucose by cellular respiration, in which glucose is broken down into water and carbon dioxide in the presence of oxygen. Virtually all organisms carry out forms of glucose oxidation; details are covered later, in chapter 7.

### Learning Outcomes Review 2.3

An ionic bond is an attraction between ions of opposite charge in an ionic compound. A covalent bond is formed when two atoms share one or more pairs of electrons. Complex biological compounds are formed in large part by atoms that can form one or more covalent bonds: C, H, O, and N. A polar covalent bond is formed by unequal sharing of electrons. Nonpolar bonds exhibit equal sharing of electrons.

- How is a polar covalent bond different from an ionic bond?

## 2.4 Water: A Vital Compound

### Learning Outcomes

1. Relate how the structure of water leads to hydrogen bonds.
2. Describe water's cohesive and adhesive properties.

Of all the common molecules, only water exists as a liquid at the relatively low temperatures that prevail on the Earth's surface. Three-fourths of the Earth is covered by liquid water



*a.* Solid

*b.* Liquid

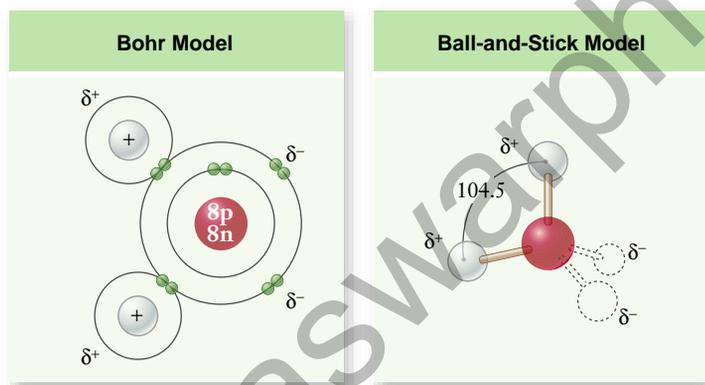
*c.* Gas

**Figure 2.10 Water takes many forms.** *a.* When water cools below  $0^{\circ}\text{C}$ , it forms beautiful crystals, familiar to us as snow and ice. *b.* Ice turns to liquid when the temperature is above  $0^{\circ}\text{C}$ . *c.* Liquid water becomes steam when the temperature rises above  $100^{\circ}\text{C}$ , as seen in this hot spring at Yellowstone National Park.

(figure 2.10). When life was beginning, water provided a medium in which other molecules could move around and interact, without being held in place by strong covalent or ionic bonds. Life evolved in water for 2 billion years before spreading to land. And even today, life is inextricably tied to water. About two-thirds of any organism's body is composed of water, and all organisms require a water-rich environment, either inside or outside it, for growth and reproduction. It is no accident that tropical rain forests are bursting with life, while dry deserts appear almost lifeless except when water becomes temporarily plentiful, such as after a rainstorm.

### Water's structure facilitates hydrogen bonding

Water has a simple molecular structure, consisting of an oxygen atom bound to two hydrogen atoms by two single covalent bonds (figure 2.11). The resulting molecule is stable: It satisfies the octet rule, has no unpaired electrons, and carries no net electrical charge.



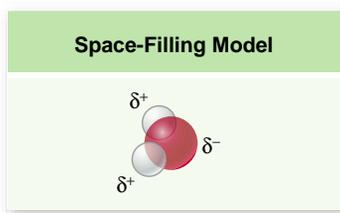
*a.*

*b.*

### Figure 2.11 Water has a simple molecular structure.

*a.* Each water molecule is composed of one oxygen atom and two hydrogen atoms. The oxygen atom shares one electron with each hydrogen atom.

*b.* The greater electronegativity of the oxygen atom makes the water molecule polar: Water carries two partial negative charges ( $\delta^{-}$ ) near the oxygen atom and two partial positive charges ( $\delta^{+}$ ), one on each hydrogen atom. *c.* Space-filling model shows what the molecule would look like if it were visible.



*c.*

The single most outstanding chemical property of water is its ability to form weak chemical associations, called **hydrogen bonds**. These bonds form between the partially negative O atoms and the partially positive H atoms of two water molecules. Although these bonds have only 5–10% of the strength of covalent bonds, they are important to DNA and protein structure, and thus responsible for much of the chemical organization of living systems.

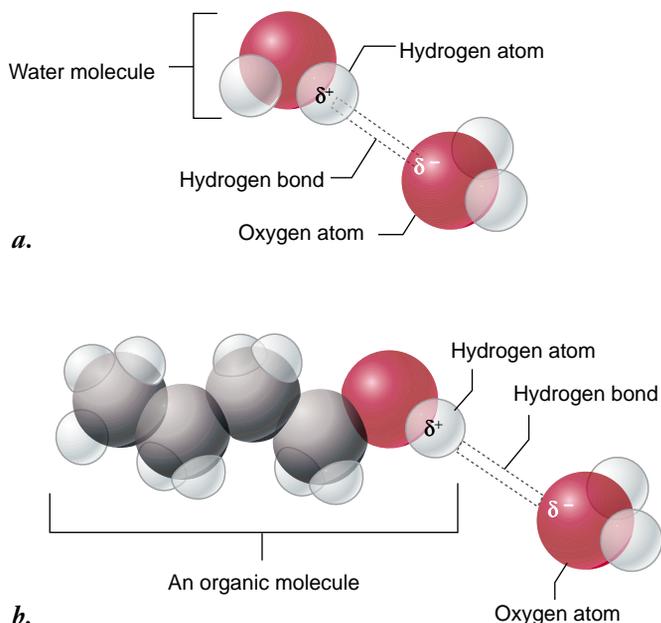
The electronegativity of O is much greater than that of H (see table 2.2), and so the bonds between these atoms are highly polar. *The polarity of water underlies water's chemistry and the chemistry of life.*

If we consider the shape of a water molecule, we see that its two covalent bonds have a partial charge at each end:  $\delta^{-}$  at the oxygen end and  $\delta^{+}$  at the hydrogen end. The most stable arrangement of these charges is a *tetrahedron* (a pyramid with a triangle as its base), in which the two negative and two positive charges are approximately equidistant from one another. The oxygen atom lies at the center of the tetrahedron, the hydrogen atoms occupy two of the apexes (corners), and the partial negative charges occupy the other two apexes (figure 2.11*b*). The bond angle between the two covalent oxygen–hydrogen bonds is  $104.5^{\circ}$ . This value is slightly less than the bond angle of a regular tetrahedron, which would be  $109.5^{\circ}$ . In water, the partial negative charges occupy more space than the partial positive regions, so the oxygen–hydrogen bond angle is slightly compressed.

### Water molecules are cohesive

The polarity of water allows water molecules to be attracted to one another: that is, water is *cohesive*. The oxygen end of each water molecule, which is  $\delta^{-}$ , is attracted to the hydrogen end, which is  $\delta^{+}$ , of other molecules. The attraction produces hydrogen bonds among water molecules (figure 2.12). Each hydrogen bond is individually very weak and transient, lasting on average only a hundred-billionth ( $10^{-11}$ ) of a second. The cumulative effects of large numbers of these bonds, however, can be enormous. Water forms an abundance of hydrogen bonds, which are responsible for many of its important physical properties (table 2.3).

Water's cohesion is responsible for its being a liquid, not a gas, at moderate temperatures. The cohesion of liquid water is also responsible for its **surface tension**. Small insects can walk on water (figure 2.13) because at the air–water interface, all the surface water molecules are hydrogen-bonded to molecules below them.



**Figure 2.12 Structure of a hydrogen bond.** *a.* Hydrogen bond between two water molecules. *b.* Hydrogen bond between an organic molecule (*n*-butanol) and water. H in *n*-butanol forms a hydrogen bond with oxygen in water. This kind of hydrogen bond is possible any time H is bound to a more electronegative atom (see table 2.2).

### Water molecules are adhesive

The polarity of water causes it to be attracted to other polar molecules as well. This attraction for other polar substances is called *adhesion*. Water adheres to any substance with which it can form hydrogen bonds. This property explains why substances containing polar molecules get “wet” when they are immersed in water, but those that are composed of nonpolar molecules (such as oils) do not.

The attraction of water to substances that have electrical charges on their surface is responsible for capillary action. If a glass tube with a narrow diameter is lowered into a beaker of water, the water will rise in the tube above the level of the water in the beaker,



**Figure 2.13 Cohesion.** Some insects, such as this water strider, literally walk on water. Because the surface tension of the water is greater than the force of one foot, the strider glides atop the surface of the water rather than sinking. The high surface tension of water is due to hydrogen bonding between water molecules.

because the adhesion of water to the glass surface, drawing it upward, is stronger than the force of gravity, pulling it downward. The narrower the tube, the greater the electrostatic forces between the water and the glass, and the higher the water rises (figure 2.14).



**Figure 2.14 Adhesion.** Capillary action causes the water within a narrow tube to rise above the surrounding water level; the adhesion of the water to the glass surface, which draws water upward, is stronger than the force of gravity, which tends to pull it down. The narrower the tube, the greater the surface area available for adhesion for a given volume of water, and the higher the water rises in the tube.

**TABLE 2.3** The Properties of Water

Property	Explanation	Example of Benefit to Life
Cohesion	Hydrogen bonds hold water molecules together.	Leaves pull water upward from the roots; seeds swell and germinate.
High specific heat	Hydrogen bonds absorb heat when they break and release heat when they form, minimizing temperature changes.	Water stabilizes the temperature of organisms and the environment.
High heat of vaporization	Many hydrogen bonds must be broken for water to evaporate.	Evaporation of water cools body surfaces.
Lower density of ice	Water molecules in an ice crystal are spaced relatively far apart because of hydrogen bonding.	Because ice is less dense than water, lakes do not freeze solid, allowing fish and other life in lakes to survive the winter.
Solubility	Polar water molecules are attracted to ions and polar compounds, making these compounds soluble.	Many kinds of molecules can move freely in cells, permitting a diverse array of chemical reactions.

### Learning Outcomes Review 2.4

Because of its polar covalent bonds, water can form hydrogen bonds with itself and with other polar molecules. Hydrogen bonding is responsible for water's cohesion, the force that holds water molecules together, and its adhesion, which is its ability to "stick" to other polar molecules. Capillary action results from both of these properties.

- If water were made of C and H instead of H and O, would it still be cohesive and adhesive?

## 2.5 Properties of Water

### Learning Outcomes

1. Describe how hydrogen bonding determines many properties of water.
2. Explain the relevance of water's unusual properties for living systems.
3. Understand the dissociation products of water.

Water moderates temperature through two properties: its high specific heat and its high heat of vaporization. Water also has the unusual property of being less dense in its solid form, ice, than as a liquid. Water acts as a solvent for polar molecules and exerts an organizing effect on nonpolar molecules. All these properties result from its polar nature.

### Water's high specific heat helps maintain temperature

The temperature of any substance is a measure of how rapidly its individual molecules are moving. In the case of water, a large input of thermal energy is required to break the many hydrogen bonds that keep individual water molecules from moving about. Therefore, water is said to have a high **specific heat**, which is defined as the amount of heat 1 g of a substance must absorb or lose to change its temperature by 1 degree Celsius ( $^{\circ}\text{C}$ ). Specific heat measures the extent to which a substance resists changing its temperature when it absorbs or loses heat. Because polar substances tend to form hydrogen bonds, the more polar it is, the higher is its specific heat. The specific heat of water (1 calorie/ $\text{g}/^{\circ}\text{C}$ ) is twice that of most carbon compounds and nine times that of iron. Only ammonia, which is more polar than water and forms very strong hydrogen bonds, has a higher specific heat than water (1.23 cal/ $\text{g}/^{\circ}\text{C}$ ). Still, only 20% of the hydrogen bonds are broken as water heats from  $0^{\circ}$  to  $100^{\circ}\text{C}$ .

Because of its high specific heat, water heats up more slowly than almost any other compound and holds its temperature longer. Because organisms have a high water content, water's high specific heat allows them to maintain a relatively constant internal temperature. The heat generated by the chemical reactions inside cells would destroy the cells if not for the absorption of this heat by the water within them.

### Water's high heat of vaporization facilitates cooling

The **heat of vaporization** is defined as the amount of energy required to change 1 g of a substance from a liquid to a gas. A considerable amount of heat energy (586 cal) is required to accomplish this change in water. As water changes from a liquid to a gas it requires energy (in the form of heat) to break its many hydrogen bonds. The evaporation of water from a surface cools that surface. Many organisms dispose of excess body heat by evaporative cooling, for example, through sweating in humans and many other vertebrates.

### Solid water is less dense than liquid water

At low temperatures, water molecules are locked into a crystal-like lattice of hydrogen bonds, forming solid ice (see figure 2.10a). Interestingly, ice is less dense than liquid water because the hydrogen bonds in ice space the water molecules relatively far apart. This unusual feature enables icebergs to float. If water did not have this property, nearly all bodies of water would be ice, with only the shallow surface melting every year. The buoyancy of ice is important ecologically because it means bodies of water freeze from the top down and not the bottom up. Because ice floats on the surface of lakes in the winter and the water beneath the ice remains liquid, fish and other animals keep from freezing.

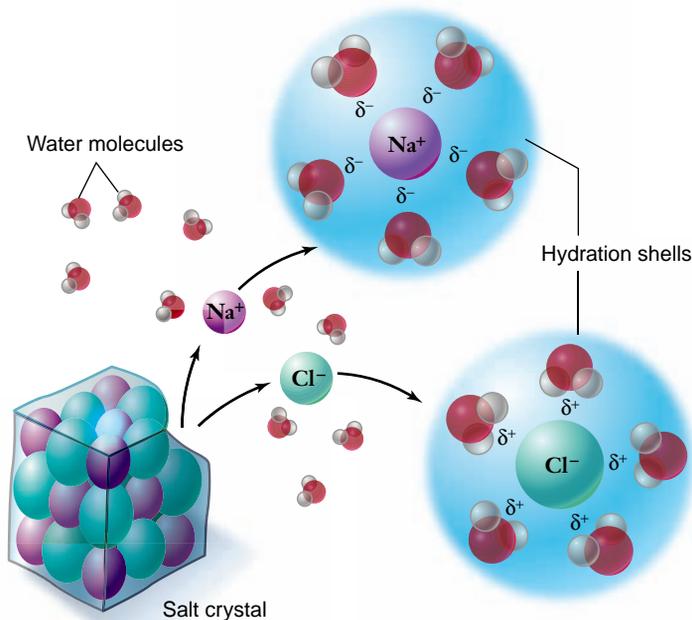
### The solvent properties of water help move ions and polar molecules

Water molecules gather closely around any substance that bears an electrical charge, whether that substance carries a full charge (ion) or a charge separation (polar molecule). For example, sucrose (table sugar) is composed of molecules that contain polar hydroxyl (OH) groups. A sugar crystal dissolves rapidly in water because water molecules can form hydrogen bonds with individual hydroxyl groups of the sucrose molecules. Therefore, sucrose is said to be *soluble* in water. Water is termed the *solvent*, and sugar is called the *solute*. Every time a sucrose molecule dissociates, or breaks away, from a solid sugar crystal, water molecules surround it in a cloud, forming a *hydration shell* that prevents it from associating with other sucrose molecules. Hydration shells also form around ions such as  $\text{Na}^+$  and  $\text{Cl}^-$  (figure 2.15).

### Water organizes nonpolar molecules

Water molecules always tend to form the maximum possible number of hydrogen bonds. When nonpolar molecules such as oils, which do not form hydrogen bonds, are placed in water, the water molecules act to exclude them. The nonpolar molecules aggregate, or clump together, thus minimizing their disruption of the hydrogen bonding of water. In effect, they shrink from contact with water, and for this reason they are referred to as **hydrophobic** (Greek *hydros*, "water," and *phobos*, "fearing"). In contrast, polar molecules, which readily form hydrogen bonds with water, are said to be **hydrophilic** ("water-loving").

The tendency of nonpolar molecules to aggregate in water is known as **hydrophobic exclusion**. By forcing the hydrophobic portions of molecules together, water causes these molecules to

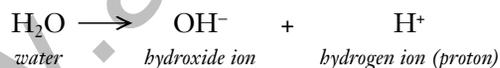


**Figure 2.15 Why salt dissolves in water.** When a crystal of table salt dissolves in water, individual  $\text{Na}^+$  and  $\text{Cl}^-$  ions break away from the salt lattice and become surrounded by water molecules. Water molecules orient around  $\text{Na}^+$  so that their partial negative poles face toward the positive  $\text{Na}^+$ ; water molecules surrounding  $\text{Cl}^-$  orient in the opposite way, with their partial positive poles facing the negative  $\text{Cl}^-$ . Surrounded by hydration shells,  $\text{Na}^+$  and  $\text{Cl}^-$  never reenter the salt lattice.

assume particular shapes. This property can also affect the structure of proteins, DNA, and biological membranes. In fact, the interaction of nonpolar molecules and water is critical to living systems.

## Water can form ions

The covalent bonds of a water molecule sometimes break spontaneously. In pure water at  $25^\circ\text{C}$ , only 1 out of every 550 million water molecules undergoes this process. When it happens, a proton (hydrogen atom nucleus) dissociates from the molecule. Because the dissociated proton lacks the negatively charged electron it was sharing, its positive charge is no longer counterbalanced, and it becomes a hydrogen ion,  $\text{H}^+$ . The rest of the dissociated water molecule, which has retained the shared electron from the covalent bond, is negatively charged and forms a hydroxide ion,  $\text{OH}^-$ . This process of spontaneous ion formation is called *ionization*:



At  $25^\circ\text{C}$ , 1 liter (L) of water contains one ten-millionth (or  $10^{-7}$ ) mole of  $\text{H}^+$  ions. A **mole** (mol) is defined as the weight of a substance in grams that corresponds to the atomic masses of all of the atoms in a molecule of that substance. In the case of  $\text{H}^+$ , the atomic mass is 1, and a mole of  $\text{H}^+$  ions would weigh 1 g. One mole of any substance always contains  $6.02 \times 10^{23}$  molecules of the substance. Therefore, the **molar concentration** of hydrogen ions in pure water, represented as  $[\text{H}^+]$ , is  $10^{-7}$  mol/L. (In reality, the  $\text{H}^+$  usually associates with another water molecule to form a hydronium ion,  $\text{H}_3\text{O}^+$ .)

## Learning Outcomes Review 2.5

Water has a high specific heat so it does not change temperature rapidly, which helps living systems maintain a near-constant temperature. Water's high heat of vaporization allows cooling by evaporation. Solid water is less dense than liquid water because the hydrogen bonds space the molecules farther apart. Polar molecules are soluble in a water solution, but water tends to exclude nonpolar molecules. Water dissociates to form  $\text{H}^+$  and  $\text{OH}^-$ .

- How does the fact that ice floats affect life in a lake?

## 2.6 Acids and Bases

### Learning Outcomes

1. Explain the nature of acids and bases, and their relationship to the pH scale.
2. Relate changes in pH to changes in  $[\text{H}^+]$ .

The concentration of hydrogen ions, and concurrently of hydroxide ions, in a solution is described by the terms *acidity* and *basicity*, respectively. Pure water, having an  $[\text{H}^+]$  of  $10^{-7}$  mol/L, is considered to be neutral, that is, neither acidic nor basic. Recall that for every  $\text{H}^+$  ion formed when water dissociates, an  $\text{OH}^-$  ion is also formed, meaning that the dissociation of water produces  $\text{H}^+$  and  $\text{OH}^-$  in equal amounts.

### The pH scale measures hydrogen ion concentration

The *pH scale* (figure 2.16) is a more convenient way to express the hydrogen ion concentration of a solution. This scale defines *pH*, which stands for “partial hydrogen,” as the negative logarithm of the hydrogen ion concentration in the solution:

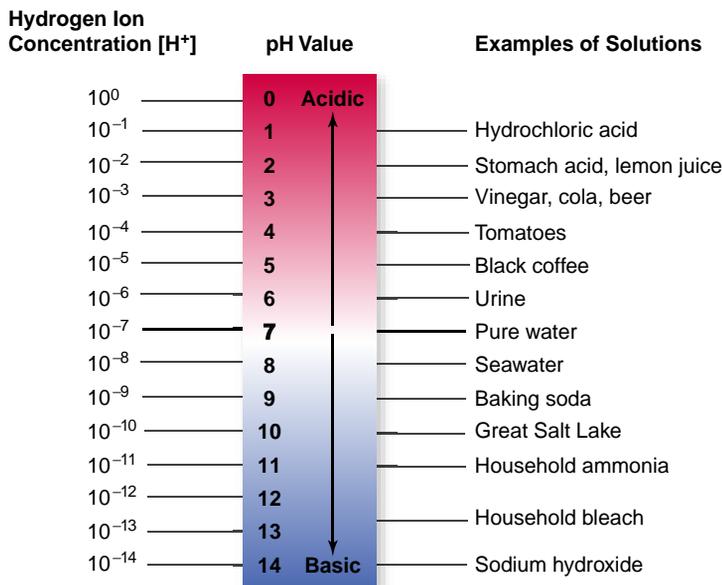
$$\text{pH} = -\log [\text{H}^+]$$

Because the logarithm of the hydrogen ion concentration is simply the exponent of the molar concentration of  $\text{H}^+$ , the pH equals the exponent times  $-1$ . For water, therefore, an  $[\text{H}^+]$  of  $10^{-7}$  mol/L corresponds to a pH value of 7. This is the neutral point—a balance between  $\text{H}^+$  and  $\text{OH}^-$ —on the pH scale. This balance occurs because the dissociation of water produces equal amounts of  $\text{H}^+$  and  $\text{OH}^-$ .

Note that, because the pH scale is *logarithmic*, a difference of 1 on the scale represents a 10-fold change in  $[\text{H}^+]$ . A solution with a pH of 4 therefore has 10 times the  $[\text{H}^+]$  of a solution with a pH of 5 and 100 times the  $[\text{H}^+]$  of a solution with a pH of 6.

### Acids

Any substance that dissociates in water to increase the  $[\text{H}^+]$  (and lower the pH) is called an **acid**. The stronger an acid is, the more hydrogen ions it produces and the lower its pH. For example, hydrochloric acid (HCl), which is abundant in your stomach, ionizes completely in water. A dilution of  $10^{-1}$  mol/L of HCl dissociates to form  $10^{-1}$  mol/L of  $\text{H}^+$ , giving the solution



**Figure 2.16 The pH scale.** The pH value of a solution indicates its concentration of hydrogen ions. Solutions with a pH less than 7 are acidic, whereas those with a pH greater than 7 are basic. The scale is logarithmic, which means that a pH change of 1 represents a 10-fold change in the concentration of hydrogen ions. Thus, lemon juice is 100 times more acidic than tomato juice, and seawater is 10 times more basic than pure water, which has a pH of 7.

a pH of 1. The pH of champagne, which bubbles because of the carbonic acid dissolved in it, is about 2.

### Bases

A substance that combines with H<sup>+</sup> when dissolved in water, and thus lowers the [H<sup>+</sup>], is called a **base**. Therefore, basic (or alkaline) solutions have pH values above 7. Very strong bases, such as sodium hydroxide (NaOH), have pH values of 12 or more. Many common cleaning substances, such as ammonia and bleach, accomplish their action because of their high pH.

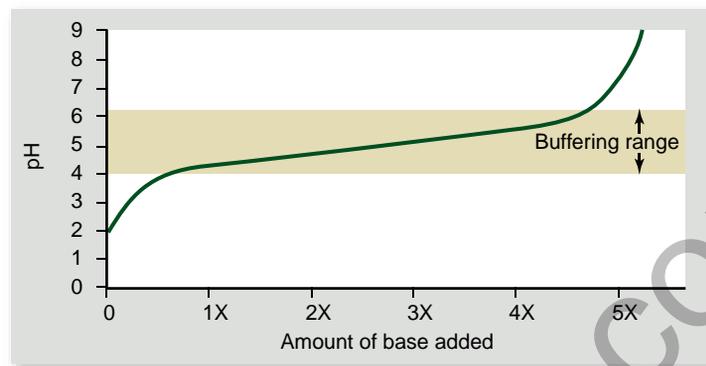
### Buffers help stabilize pH

The pH inside almost all living cells, and in the fluid surrounding cells in multicellular organisms, is fairly close to neutral, 7. Most of the enzymes in living systems are extremely sensitive to pH. Often even a small change in pH will alter their shape, thereby disrupting their activities. For this reason, it is important that a cell maintain a constant pH level.

But the chemical reactions of life constantly produce acids and bases within cells. Furthermore, many animals eat substances that are acidic or basic. Cola drinks, for example, are moderately strong (although dilute) acidic solutions. Despite such variations in the concentrations of H<sup>+</sup> and OH<sup>-</sup>, the pH of an organism is kept at a relatively constant level by buffers (figure 2.17).

A **buffer** is a substance that resists changes in pH. Buffers act by releasing hydrogen ions when a base is added and absorbing hydrogen ions when acid is added, with the overall effect of keeping [H<sup>+</sup>] relatively constant.

Within organisms, most buffers consist of pairs of substances, one an acid and the other a base. The key buffer in human blood is an acid–base pair consisting of carbonic acid



**Figure 2.17 Buffers minimize changes in pH.** Adding a base to a solution neutralizes some of the acid present, and so raises the pH. Thus, as the curve moves to the right, reflecting more and more base, it also rises to higher pH values. A buffer makes the curve rise or fall very slowly over a portion of the pH scale, called the “buffering range” of that buffer.

### Inquiry question

**?** For this buffer, adding base raises pH more rapidly below pH 4 than above it. What might account for this behavior?

(acid) and bicarbonate (base). These two substances interact in a pair of reversible reactions. First, carbon dioxide (CO<sub>2</sub>) and H<sub>2</sub>O join to form carbonic acid (H<sub>2</sub>CO<sub>3</sub>), which in a second reaction dissociates to yield bicarbonate ion (HCO<sub>3</sub><sup>-</sup>) and H<sup>+</sup>.

If some acid or other substance adds H<sup>+</sup> to the blood, the HCO<sub>3</sub><sup>-</sup> acts as a base and removes the excess H<sup>+</sup> by forming H<sub>2</sub>CO<sub>3</sub>. Similarly, if a basic substance removes H<sup>+</sup> from the blood, H<sub>2</sub>CO<sub>3</sub> dissociates, releasing more H<sup>+</sup> into the blood. The forward and reverse reactions that interconvert H<sub>2</sub>CO<sub>3</sub> and HCO<sub>3</sub><sup>-</sup> thus stabilize the blood’s pH.



The reaction of carbon dioxide and water to form carbonic acid is a crucial one because it permits carbon, essential to life, to enter water from the air. The Earth’s oceans are rich in carbon because of the reaction of carbon dioxide with water.

In a condition called blood acidosis, human blood, which normally has a pH of about 7.4, drops to a pH of about 7.1. This condition is fatal if not treated immediately. The reverse condition, blood alkalosis, involves an increase in blood pH of a similar magnitude and is just as serious.

### Learning Outcomes Review 2.6

Acid solutions have a high [H<sup>+</sup>], and basic solutions have a low [H<sup>+</sup>] (and therefore a high [OH<sup>-</sup>]). The pH of a solution is the negative logarithm of its [H<sup>+</sup>]. Low pH values indicate acids, and high pH values indicate bases. Even small changes in pH can be harmful to life. Buffer systems in organisms help to maintain pH within a narrow range.

■ A change of 2 pH units indicates what change in [H<sup>+</sup>]?



## Chapter Review

### 2.1 The Nature of Atoms

*All matter is composed of atoms (see figure 2.3).*

**Atomic structure includes a central nucleus and orbiting electrons.**

Electrically neutral atoms have the same number of protons as electrons. Atoms that gain or lose electrons are called ions.

Each element is defined by its atomic number, the number of protons in the nucleus. Atomic mass is the sum of the mass of protons and neutrons in an atom. Isotopes are forms of a single element with different numbers of neutrons, and thus different atomic mass. Radioactive isotopes are unstable.

**Electrons determine the chemical behavior of atoms.**

The potential energy of electrons increases as distance from the nucleus increases. Electron orbitals are probability distributions. *s*-orbitals are spherical; other orbitals have different shapes, such as the dumbbell-shaped *p*-orbitals.

**Atoms contain discrete energy levels.**

Energy levels correspond to quanta (sing. quantum) of energy, a “ladder” of energy levels that an electron may have.

The loss of electrons from an atom is called oxidation. The gain of electrons is called reduction. Electrons can be transferred from one atom to another in coupled redox reactions.

### 2.2 Elements Found in Living Systems

*The periodic table displays elements according to atomic number and properties.*

Atoms tend to establish completely full outer energy levels (the octet rule). Elements with filled outermost orbitals are inert.

Ninety elements occur naturally in the Earth’s crust. Twelve of these elements are found in living organisms in greater than trace amounts: C, H, O, N, P, S, Na, K, Ca, Mg, Fe, and Cl.

Compounds of carbon are called organic compounds. The majority of molecules in living systems are composed of C bound to H, O, and N.

### 2.3 The Nature of Chemical Bonds

Molecules contain two or more atoms joined by chemical bonds. Compounds contain two or more different elements.

**Ionic bonds form crystals.**

Ions with opposite electrical charges form ionic bonds, such as NaCl (see figure 2.9*b*).

**Covalent bonds build stable molecules.**

A molecule formed by a covalent bond is stable because it has no net charge, the octet rule is satisfied, and it has no unpaired electrons. Covalent bonds may be single, double, or triple, depending on the number of pairs of electrons shared. Nonpolar covalent bonds involve equal sharing of electrons between atoms. Polar covalent bonds involve unequal sharing of electrons.

**Chemical reactions alter bonds.**

Temperature, reactant concentration, and the presence of catalysts affect reaction rates. Most biological reactions are reversible, such as the conversion of carbon dioxide and water into carbohydrates.

### 2.4 Water: A Vital Compound

**Water’s structure facilitates hydrogen bonding.**

Hydrogen bonds are weak interactions between a partially positive H in one molecule and a partially negative O in another molecule (see figure 2.11).

**Water molecules are cohesive.**

Cohesion is the tendency of water molecules to adhere to one another due to hydrogen bonding. The cohesion of water is responsible for its surface tension.

**Water molecules are adhesive.**

Adhesion occurs when water molecules adhere to other polar molecules. Capillary action results from water’s adhesion to the sides of narrow tubes, combined with its cohesion.

### 2.5 Properties of Water

**Water’s high specific heat helps maintain temperature.**

The specific heat of water is high because it takes a considerable amount of energy to disrupt hydrogen bonds.

**Water’s high heat of vaporization facilitates cooling.**

Breaking hydrogen bonds to turn liquid water into vapor takes a lot of energy. Many organisms lose excess heat through evaporative cooling, such as sweating.

**Solid water is less dense than liquid water.**

Hydrogen bonds are spaced farther apart in the solid phase of water than in the liquid phase. As a result, ice floats.

**The solvent properties of water help move ions and polar molecules.**

Water’s polarity makes it a good solvent for polar substances and ions. Polar molecules or portions of molecules are attracted to water (hydrophilic). Molecules that are nonpolar are repelled by water (hydrophobic). Water makes nonpolar molecules clump together.

**Water organizes nonpolar molecules.**

Nonpolar molecules will aggregate to avoid water. This maximizes the hydrogen bonds that water can make. This hydrophobic exclusion can affect the structure of DNA, proteins and biological membranes.

**Water can form ions.**

Water dissociates into  $H^+$  and  $OH^-$ . The concentration of  $H^+$ , shown as  $[H^+]$ , in pure water is  $10^{-7}$  mol/L.

### 2.6 Acids and Bases (see figure 2.16)

**The pH scale measures hydrogen ion concentration.**

pH is defined as the negative logarithm of  $[H^+]$ . Pure water has a pH of 7. A difference of 1 pH unit means a 10-fold change in  $[H^+]$ .

Acids have a greater  $[H^+]$  and therefore a lower pH; bases have a lower  $[H^+]$  and therefore a higher pH.

**Buffers help stabilize pH.**

Carbon dioxide and water react reversibly to form carbonic acid. A buffer resists changes in pH by absorbing or releasing  $H^+$ . The key buffer in the human blood is the carbonic acid/bicarbonate pair.

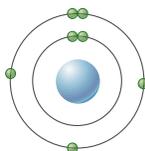
## Review Questions

### UNDERSTAND

- The property that distinguishes an atom of one element (carbon, for example) from an atom of another element (oxygen, for example) is
  - the number of electrons.
  - the number of protons.
  - the number of neutrons.
  - the combined number of protons and neutrons.
- If an atom has one valence electron, that is, a single electron in its outer energy level, it will most likely form
  - one polar, covalent bond.
  - two nonpolar, covalent bonds.
  - two covalent bonds.
  - an ionic bond.
- An atom with a net positive charge must have more
  - protons than neutrons.
  - protons than electrons.
  - electrons than neutrons.
  - electrons than protons.
- The isotopes carbon-12 and carbon-14 differ in
  - the number of neutrons.
  - the number of protons.
  - the number of electrons.
  - both b and c.
- Which of the following is NOT a property of the elements most commonly found in living organisms?
  - The elements have a low atomic mass.
  - The elements have an atomic number less than 21.
  - The elements possess eight electrons in their outer energy level.
  - The elements are lacking one or more electrons from their outer energy level.
- Ionic bonds arise from
  - shared valence electrons.
  - attractions between valence electrons.
  - charge attractions between valence electrons.
  - attractions between ions of opposite charge.
- A substance with a high concentration of hydrogen ions
  - is called a base.
  - is called an acid.
  - has a high pH.
  - both b and c.

### APPLY

- Using the periodic table on page 22, which of the following atoms would you predict could form a positively charged ion (cation)?
  - Fluorine (F)
  - Neon (Ne)
  - Potassium (K)
  - Sulfur (S)
- Refer to the element pictured. How many covalent bonds could this atom form?
  - Two
  - Three
  - Four
  - None



- A molecule with polar covalent bonds would
  - be soluble in water.
  - not be soluble in water.
  - contain atoms with very similar electronegativity.
  - contain atoms that have gained or lost electrons.
- Hydrogen bonds are formed
  - between any molecules that contain hydrogen.
  - only between water molecules.
  - when hydrogen is part of a polar bond.
  - when two atoms of hydrogen share an electron.
- Which of the following properties of water is NOT a consequence of its ability to form hydrogen bonds?
  - Cohesiveness
  - High specific heat
  - Ability to function as a solvent
  - Neutral pH
- The decay of radioactive isotopes involves changes to the nucleus of atoms. Explain how this differs from the changes in atoms that occur during chemical reactions.

### SYNTHESIZE

- Elements that form ions are important for a range of biological processes. You have learned something about the cations sodium ( $\text{Na}^+$ ), calcium ( $\text{Ca}^{2+}$ ) and potassium ( $\text{K}^+$ ) in this chapter. Use your knowledge of the definition of a cation to identify other examples from the periodic table.
- A popular theme in science fiction literature has been the idea of silicon-based life-forms in contrast to our carbon-based life. Evaluate the possibility of silicon-based life based on the chemical structure and potential for chemical bonding of a silicon atom.
- Recent efforts by NASA to search for signs of life on Mars have focused on the search for evidence of liquid water rather than looking directly for biological organisms (living or fossilized). Use your knowledge of the influence of water on life on Earth to construct an argument justifying this approach.

### ONLINE RESOURCE

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Understand, Apply, and Synthesize—enhance your study with animations that bring concepts to life and practice tests to assess your understanding. Your instructor may also recommend the interactive eBook, individualized learning tools, and more.

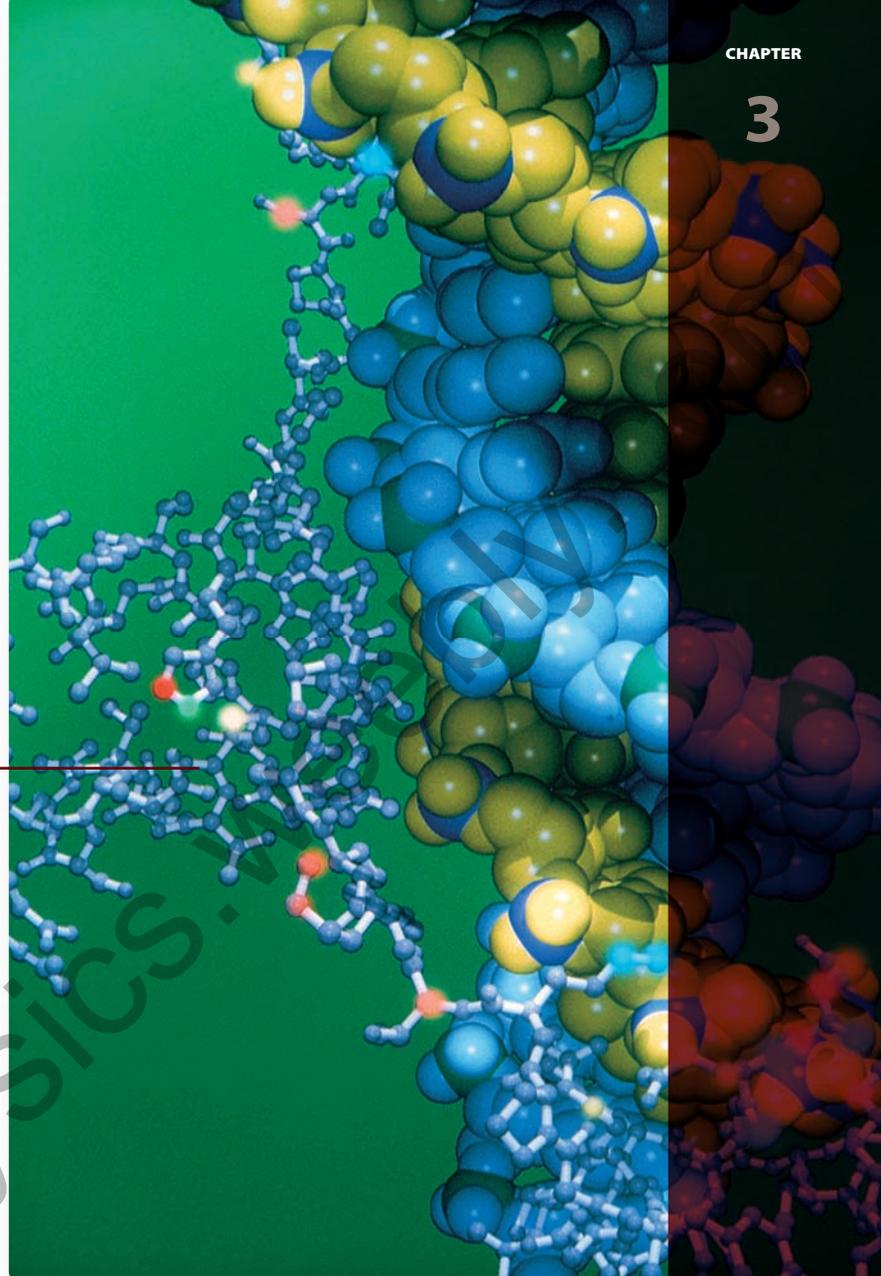
# Chapter 3

## The Chemical Building Blocks of Life

### Chapter Outline

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- 3.1 Carbon: The Framework of Biological Molecules
- 3.2 Carbohydrates: Energy Storage and Structural Molecules
- 3.3 Nucleic Acids: Information Molecules
- 3.4 Proteins: Molecules with Diverse Structures and Functions
- 3.5 Lipids: Hydrophobic Molecules



### Introduction

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A cup of water contains more molecules than there are stars in the sky. But many molecules are much larger than water molecules. Many thousands of distinct biological molecules are long chains made of thousands or even billions of atoms. These enormous assemblages, which are almost always synthesized by living things, are macromolecules. As you may know, biological macromolecules can be divided into four categories: carbohydrates, nucleic acids, proteins, and lipids, and they are the basic chemical building blocks from which all organisms are composed.

We take the existence of these classes of macromolecules for granted now, but as late as the 19th century many theories of “vital forces” were associated with living systems. One such theory held that cells contained a substance, protoplasm, that was responsible for the chemical reactions in living systems. Any disruption of cells was thought to disturb the protoplasm. Such a view makes studying the chemical reactions of cells in the lab (*in vitro*) impossible. The demonstration of fermentation in a cell-free system marked the beginning of modern biochemistry (figure 3.1). This approach involves studying biological molecules outside of cells to infer their role inside cells. Because these biological macromolecules all involve carbon-containing compounds, we begin with a brief summary of carbon and its chemistry.

## SCIENTIFIC THINKING

**Hypothesis:** Chemical reactions, such as the fermentation reaction in yeast, are controlled by enzymes and do not require living cells.

**Prediction:** If yeast cells are broken open, these enzymes should function outside of the cell.

**Test:** Yeast is mixed with quartz sand and diatomaceous earth and then ground in a mortar and pestle. The resulting paste is wrapped in canvas and subjected to 400–500 atm pressure in a press. Fermentable and nonfermentable substrates are added to the resulting fluid, with fermentation being measured by the production of  $\text{CO}_2$ .



**Result:** When a fermentable substrate (cane sugar, glucose) is used,  $\text{CO}_2$  is produced, when a nonfermentable substrate (lactose, mannose) is used, no  $\text{CO}_2$  is produced. In addition, visual inspection of the fluid shows no visible yeast cells.

**Conclusion:** The hypothesis is supported. The fermentation reaction can occur in the absence of live yeast.

**Historical Significance:** Although this is not precisely the intent of the original experiment, it represents the first use of a cell-free system. Such systems allow for the study of biochemical reactions *in vitro* and the purification of proteins involved. We now know that the “fermentation reaction” is actually a complex series of reactions. Would such a series of reactions be your first choice for this kind of demonstration?

**Figure 3.1** The demonstration of cell-free fermentation. Eduard Buchner’s (1860–1917) demonstration of fermentation by fluid produced from yeast, but not containing any live cells both argued against the protoplasm theory and provided a method for future biochemists to examine the chemistry of life outside of cells.

### 3.1

## Carbon: The Framework of Biological Molecules

### Learning Outcomes

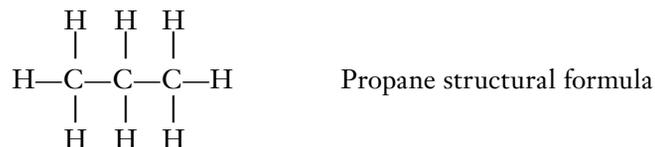
1. Describe the relationship between functional groups and macromolecules.
2. Recognize the different kinds of isomers.
3. List the different kinds of biological macromolecules.

In chapter 2, we reviewed the basics of chemistry. Biological systems obey all the laws of chemistry. Thus, chemistry forms the basis of living systems.

The framework of biological molecules consists predominantly of carbon atoms bonded to other carbon atoms or to atoms of oxygen, nitrogen, sulfur, phosphorus, or hydrogen. Because carbon atoms can form up to four covalent bonds, molecules containing carbon can form straight chains, branches, or even rings, balls, tubes, and coils.

Molecules consisting only of carbon and hydrogen are called *hydrocarbons*. Because carbon–hydrogen covalent bonds store considerable energy, hydrocarbons make good fuels. Gasoline, for example, is rich in hydrocarbons, and propane gas, an-

other hydrocarbon, consists of a chain of three carbon atoms, with eight hydrogen atoms bound to it. The chemical formula for propane is  $\text{C}_3\text{H}_8$ . Its structural formula is



Theoretically speaking, the length of a chain of carbon atoms is unlimited. As described in the rest of this chapter, the four main types of biological molecules often consist of huge chains of carbon-containing compounds.

### Functional groups account for differences in molecular properties

Carbon and hydrogen atoms both have very similar electronegativities. Electrons in  $\text{C}-\text{C}$  and  $\text{C}-\text{H}$  bonds are therefore evenly distributed, with no significant differences in charge over the molecular surface. For this reason, hydrocarbons are nonpolar. Most biological molecules produced by cells, however, also contain other atoms. Because these other atoms frequently have different electronegativities, molecules containing them exhibit regions of partial positive or negative charge. They are polar. These molecules can be thought of as a  $\text{C}-\text{H}$

core to which specific molecular groups, called **functional groups**, are attached. One such common functional group is  $\text{—OH}$ , called a *hydroxyl group*.

Functional groups have definite chemical properties that they retain no matter where they occur. Both the hydroxyl and carbonyl ( $\text{C=O}$ ) groups, for example, are polar because of the

electronegativity of the oxygen atoms (see chapter 2). Other common functional groups are the acidic carboxyl ( $\text{COOH}$ ), phosphate ( $\text{PO}_4$ ), and the basic amino ( $\text{NH}_2$ ) group. Many of these functional groups can also participate in hydrogen bonding. Hydrogen bond donors and acceptors can be predicted based on their electronegativities shown in table 2.2. Figure 3.2 illustrates these biologically important functional groups and lists the macromolecules in which they are found.

Functional Group	Structural Formula	Example	Found In
Hydroxyl	$\text{—OH}$	$\begin{array}{c} \text{H} \quad \text{H} \\   \quad   \\ \text{H—C—C—OH} \\   \quad   \\ \text{H} \quad \text{H} \end{array}$ Ethanol	carbohydrates, proteins, nucleic acids, lipids
Carbonyl	$\begin{array}{c} \text{O} \\    \\ \text{—C—} \end{array}$	$\begin{array}{c} \text{H} \quad \text{O} \\   \quad    \\ \text{H—C—C—H} \\   \\ \text{H} \end{array}$ Acetaldehyde	carbohydrates, nucleic acids
Carboxyl	$\begin{array}{c} \text{O} \\ // \\ \text{—C} \\ \backslash \\ \text{OH} \end{array}$	$\begin{array}{c} \text{H} \quad \text{O} \\   \quad    \\ \text{H—C—C—OH} \\   \\ \text{H} \end{array}$ Acetic acid	proteins, lipids
Amino	$\begin{array}{c} \text{H} \\   \\ \text{—N—} \\   \\ \text{H} \end{array}$	$\begin{array}{c} \text{O} \quad \text{H} \\    \quad   \\ \text{HO—C—C—N—H} \\ \quad \quad   \\ \quad \quad \text{CH}_3 \end{array}$ Alanine	proteins, nucleic acids
Sulfhydryl	$\text{—S—H}$	$\begin{array}{c} \text{COOH} \\   \\ \text{H—C—CH}_2\text{—S—H} \\   \\ \text{NH}_2 \end{array}$ Cysteine	proteins
Phosphate	$\begin{array}{c} \text{O}^- \\   \\ \text{—O—P—O}^- \\    \\ \text{O} \end{array}$	$\begin{array}{c} \text{OH} \quad \text{OH} \quad \text{H} \quad \text{O} \\   \quad   \quad   \quad    \\ \text{H—C—C—C—O—P—O}^- \\   \quad   \quad   \quad   \\ \text{H} \quad \text{H} \quad \text{H} \quad \text{O}^- \end{array}$ Glycerol phosphate	nucleic acids
Methyl	$\begin{array}{c} \text{H} \\   \\ \text{—C—H} \\   \\ \text{H} \end{array}$	$\begin{array}{c} \text{O} \quad \text{H} \\    \quad   \\ \text{HO—C—C—NH}_2 \\   \\ \text{H—C—H} \\   \\ \text{H} \end{array}$ Alanine	proteins

**Figure 3.2** The primary functional chemical groups.

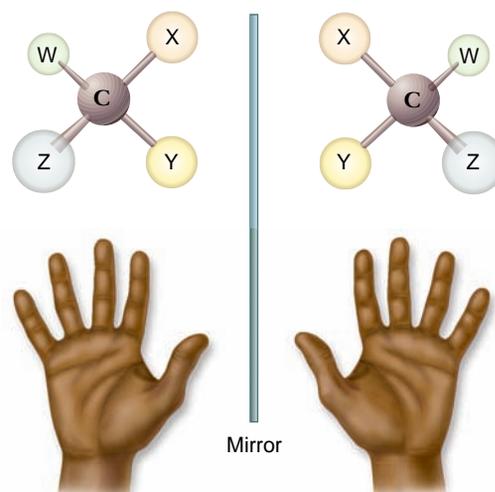
These groups tend to act as units during chemical reactions and give specific chemical properties to the molecules that possess them. Amino groups, for example, make a molecule more basic, and carboxyl groups make a molecule more acidic. These functional groups are also not limited to the examples in the “Found In” column but are widely distributed in biological molecules.

### Isomers have the same molecular formulas but different structures

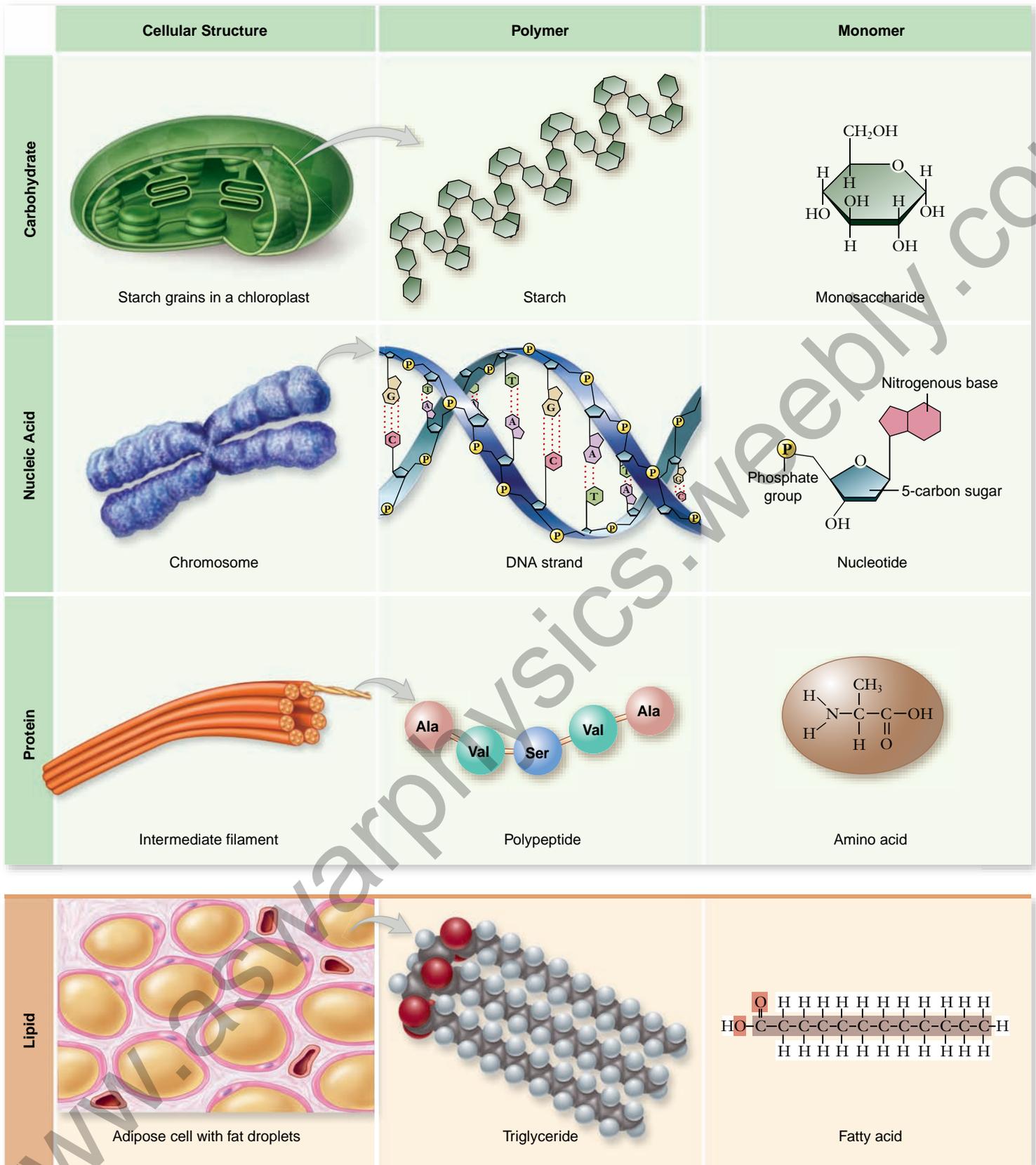
Organic molecules having the same molecular or empirical formula can exist in different forms called **isomers**. If there are differences in the actual structure of their carbon skeleton, we call them *structural isomers*. Later you will see that glucose and fructose are structural isomers of  $\text{C}_6\text{H}_{12}\text{O}_6$ . Another form of isomers, called *stereoisomers*, have the same carbon skeleton but differ in how the groups attached to this skeleton are arranged in space.

Enzymes in biological systems usually recognize only a single, specific stereoisomer. A subcategory of stereoisomers, called *enantiomers*, are actually mirror images of each other. A molecule that has mirror-image versions is called a *chiral* molecule. When carbon is bound to four different molecules, this inherent asymmetry exists (figure 3.3).

Chiral compounds are characterized by their effect on polarized light. Polarized light has a single plane, and chiral molecules rotate this plane either to the right (Latin, *dextro*) or left (Latin, *levo*). We therefore call the two chiral forms *D* for *dextrorotatory* and *L* for *levorotatory*. Living systems tend to produce only a single enantiomer of the two possible forms; for example, in most organisms we find primarily D-sugars and L-amino acids.



**Figure 3.3** Chiral molecules. When carbon is bound to four different groups, the resulting molecule is said to be chiral (from Greek *cheir*, meaning “hand.”). A chiral molecule will have stereoisomers that are mirror images. The two molecules shown have the same four groups but cannot be superimposed, much like your two hands cannot be superimposed but must be flipped to match. These types of stereoisomers are called *enantiomers*.



**Figure 3.4 Polymer macromolecules.** The four major biological macromolecules are shown. Carbohydrates, nucleic acids, and proteins all form polymers and are shown with the monomers used to make them. Lipids do not fit this simple monomer–polymer relationship, however, because they are constructed from glycerol and fatty acids. All four types of macromolecules are also shown in their cellular context.

TABLE 3.1		Macromolecules	
Macromolecule	Subunit	Function	Example
<b>C A R B O H Y D R A T E S</b>			
Starch, glycogen	Glucose	Energy storage	Potatoes
Cellulose	Glucose	Structural support in plant cell walls	Paper; strings of celery
Chitin	Modified glucose	Structural support	Crab shells
<b>N U C L E I C A C I D S</b>			
DNA	Nucleotides	Encodes genes	Chromosomes
RNA	Nucleotides	Needed for gene expression	Messenger RNA
<b>P R O T E I N S</b>			
Functional	Amino acids	Catalysis; transport	Hemoglobin
Structural	Amino acids	Support	Hair; silk
<b>L I P I D S</b>			
Fats	Glycerol and three fatty acids	Energy storage	Butter; corn oil; soap
Phospholipids	Glycerol, two fatty acids, phosphate, and polar R groups	Cell membranes	Phosphatidylcholine
Prostaglandins	Five-carbon rings with two nonpolar tails	Chemical messengers	Prostaglandin E (PGE)
Steroids	Four fused carbon rings	Membranes; hormones	Cholesterol; estrogen
Terpenes	Long carbon chains	Pigments; structural support	Carotene; rubber

## Biological macromolecules include carbohydrates, nucleic acids, proteins, and lipids

Remember that biological macromolecules are traditionally grouped into carbohydrates, nucleic acids, proteins, and lipids (table 3.1). In many cases, these macromolecules are polymers. A **polymer** is a long molecule built by linking together a large number of small, similar chemical subunits called **monomers**. They are like railroad cars coupled to form a train. The nature of a polymer is determined by the monomers used to build the polymer. Here are some examples. Complex carbohydrates such as starch are polymers composed of simple ring-shaped sugars. Nucleic acids (DNA and RNA) are polymers of nucleotides (figure 3.4). Proteins are polymers of amino acids, and lipids are polymers of fatty acids (see figure 3.4). These long chains are built via chemical reactions termed *dehydration reactions* and are broken down by *hydrolysis reactions*.

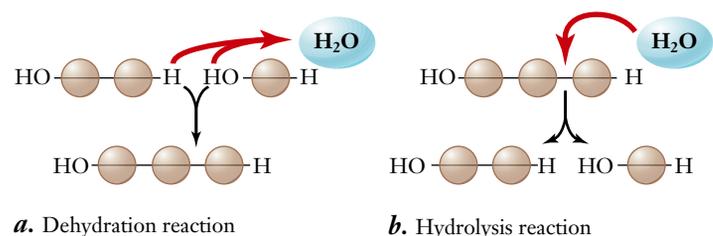
### The dehydration reaction

Despite the differences between monomers of these major polymers, the basic chemistry of their synthesis is similar: To form a covalent bond between two monomers, an —OH group is removed from one monomer, and a hydrogen atom (H) is removed from the other (figure 3.5*a*). For example, this simple chemistry is the same for linking amino acids together to make a protein or assembling glucose units together to make starch. This reaction is also used to link fatty acids to glycerol in lipids. This chemical reaction is called condensation, or a **dehydration reaction**, because the removal of —OH and —H is the same as the removal

of a molecule of water ( $H_2O$ ). For every subunit added to a macromolecule, one water molecule is removed. These and other biochemical reactions require that the reacting substances are held close together and that the correct chemical bonds are stressed and broken. This process of positioning and stressing, termed *catalysis*, is carried out within cells by enzymes.

### The hydrolysis reaction

Cells disassemble macromolecules into their constituent subunits through reactions that are the reverse of dehydration—a molecule of water is added instead of removed (figure 3.5*b*). In this process, called **hydrolysis**, a hydrogen atom is attached to one subunit and a hydroxyl group to the other, breaking a specific covalent bond in the macromolecule.



## Figure 3.5 Making and breaking macromolecules.

**a.** Biological macromolecules are polymers formed by linking monomers together through dehydration reactions. This process releases a water molecule for every bond formed. **b.** Breaking the bond between subunits involves hydrolysis, which reverses the loss of a water molecule by dehydration.

### Learning Outcomes Review 3.1

Functional groups account for differences in chemical properties in organic molecules. Isomers are compounds with the same empirical formula but different structures. This difference may affect biological function. Macromolecules are polymers consisting of long chains of similar subunits that are joined by dehydration reactions and are broken down by hydrolysis reactions.

- What is the relationship between dehydration and hydrolysis?

## 3.2 Carbohydrates: Energy Storage and Structural Molecules

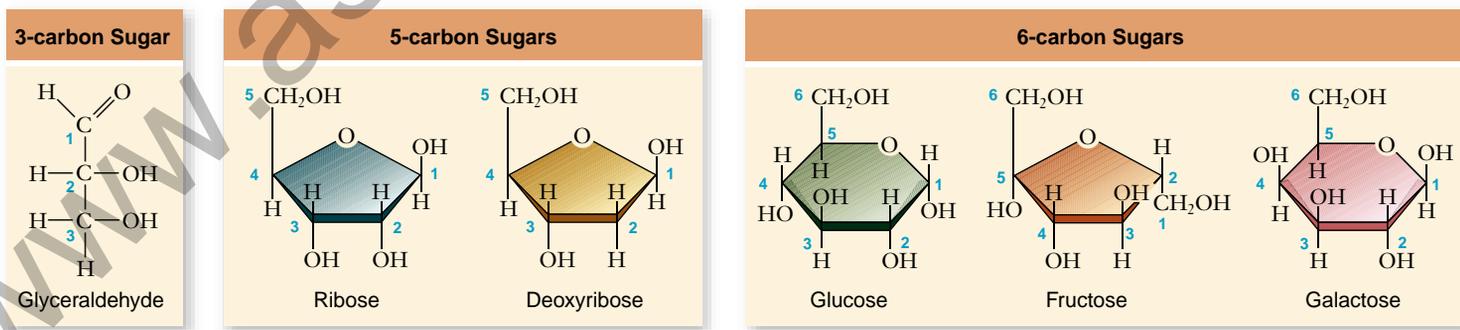
### Learning Outcomes

1. Describe the structure of a sugar.
2. Name the different forms of carbohydrate molecules.
3. Relate the structure of polysaccharides to their functions

### Monosaccharides are simple sugars

Carbohydrates are a loosely defined group of molecules that all contain carbon, hydrogen, and oxygen in the molar ratio 1:2:1. Their empirical formula (which lists the number of atoms in the molecule with subscripts) is  $(\text{CH}_2\text{O})_n$ , where  $n$  is the number of carbon atoms. Because they contain many carbon-hydrogen (C—H) bonds, which release energy when oxidation occurs, carbohydrates are well suited for energy storage. Sugars are among the most important energy-storage molecules, and they exist in several different forms.

The simplest of the carbohydrates are the **monosaccharides** (Greek *mono*, “single,” and Latin *saccharum*, “sugar”). Simple sugars contain as few as three carbon atoms, but those that play the central role in energy storage have six (figure 3.6). The empirical formula of six-carbon sugars is:



**Figure 3.6 Monosaccharides.** Monosaccharides, or simple sugars, can contain as few as three carbon atoms and are often used as building blocks to form larger molecules. The five-carbon sugars ribose and deoxyribose are components of nucleic acids (see figure 3.15). The carbons are conventionally numbered from the more oxidized end.

Six-carbon sugars can exist in a straight-chain form, but dissolved in water (an aqueous environment) they almost always form rings.

The most important of the six-carbon monosaccharides for energy storage is glucose, which you first encountered in the examples of chemical reactions in chapter 2. Glucose has seven energy-storing C—H bonds (figure 3.7). Depending on the orientation of the carbonyl group (C=O) when the ring is closed, glucose can exist in two different forms: alpha ( $\alpha$ ) or beta ( $\beta$ ).

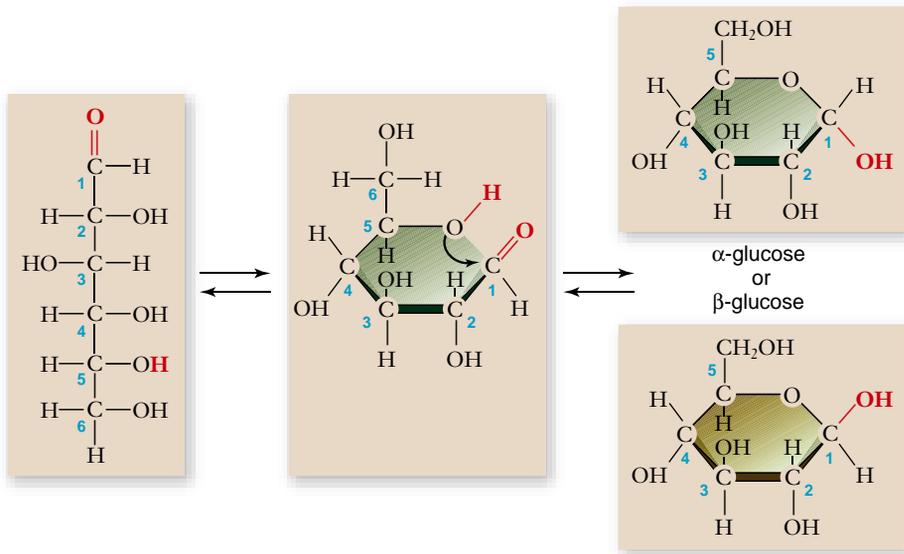
### Sugar isomers have structural differences

Glucose is not the only sugar with the formula  $\text{C}_6\text{H}_{12}\text{O}_6$ . Both structural isomers and stereoisomers of this simple six-carbon skeleton exist in nature. Fructose is a structural isomer that differs in the position of the carbonyl carbon (C=O); galactose is a stereoisomer that differs in the position of —OH and —H groups relative to the ring (figure 3.8). These differences often account for substantial functional differences between the isomers. Your taste buds can discern them: Fructose tastes much sweeter than glucose, despite the fact that both sugars have identical chemical composition. Enzymes that act on different sugars can distinguish both the structural and stereoisomers of this basic six-carbon skeleton. The different stereoisomers of glucose are also important in the polymers that can be made using glucose as a monomer, as you will see later in this chapter.

### Disaccharides serve as transport molecules in plants and provide nutrition in animals

Most organisms transport sugars within their bodies. In humans, the glucose that circulates in the blood does so as a simple monosaccharide. In plants and many other organisms, however, glucose is converted into a transport form before it is moved from place to place within the organism. In such a form, it is less readily metabolized during transport.

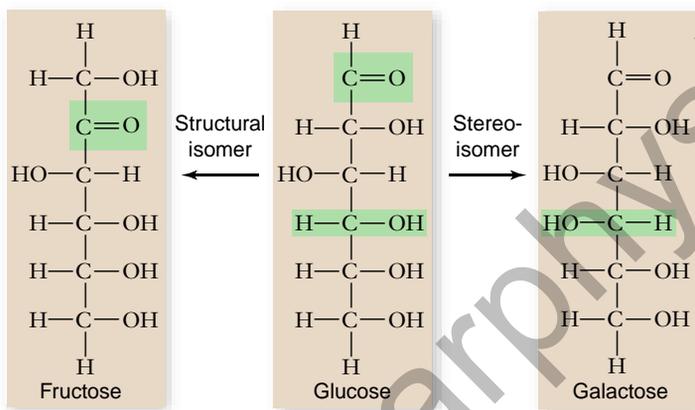
Transport forms of sugars are commonly made by linking two monosaccharides together to form a **disaccharide** (Greek *di*, “two”). Disaccharides serve as effective reservoirs of glucose because the enzymes that normally use glucose in the organism cannot break the bond linking the two monosaccharide subunits. Enzymes that can do so are typically present only in the tissue that uses glucose.



**Figure 3.7 Structure of the glucose molecule.** Glucose is a linear, six-carbon molecule that forms a six-membered ring in solution. Ring closure occurs such that two forms can result:  $\alpha$ -glucose and  $\beta$ -glucose. These structures differ only in the position of the  $\text{—OH}$  bound to carbon 1. The structure of the ring can be represented in many ways; shown here are the most common, with the carbons conventionally numbered (in blue) so that the forms can be compared easily. The heavy lines in the ring structures represent portions of the molecule that are projecting out of the page toward you.

Transport forms differ depending on which monosaccharides are linked to form the disaccharide. Glucose forms transport disaccharides with itself and with many other monosaccharides, including fructose and galactose. When glucose forms a disaccharide with the structural isomer fructose, the resulting disaccharide is *sucrose*, or table sugar (figure 3.9*a*). Sucrose is the form most plants use to transport glucose and is the sugar that most humans and other animals eat. Sugarcane and sugar beets are rich in sucrose.

When glucose is linked to the stereoisomer galactose, the resulting disaccharide is *lactose*, or milk sugar. Many mammals supply energy to their young in the form of lactose. Adults often have greatly reduced levels of lactase, the enzyme required to cleave lactose into its two monosaccharide components, and thus they cannot metabolize lactose efficiently. This can result in lactose intolerance in humans. Most of the energy that is channeled into lactose production is therefore reserved for offspring. For this reason, lactose as an energy source is primarily for offspring in mammals.



**Figure 3.8 Isomers and stereoisomers.** Glucose, fructose, and galactose are isomers with the empirical formula  $\text{C}_6\text{H}_{12}\text{O}_6$ . A structural isomer of glucose, such as fructose, has identical chemical groups bonded to different carbon atoms. Notice that this results in a five-membered ring in solution (see figure 3.6). A stereoisomer of glucose, such as galactose, has identical chemical groups bonded to the same carbon atoms but in different orientations (the  $\text{—OH}$  at carbon 4).

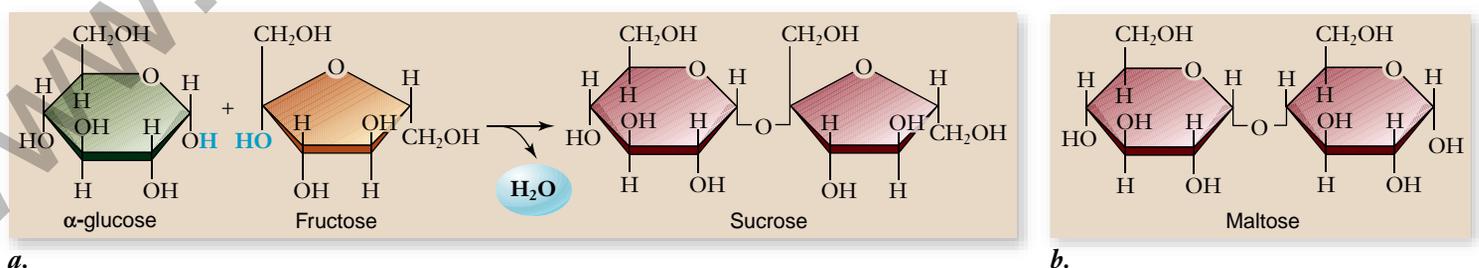
## Polysaccharides provide energy storage and structural components

**Polysaccharides** are longer polymers made up of monosaccharides that have been joined through dehydration reactions. **Starch**, a storage polysaccharide, consists entirely of  $\alpha$ -glucose molecules linked in long chains. **Cellulose**, a structural polysaccharide, also consists of glucose molecules linked in chains, but these molecules are  $\beta$ -glucose. Because starch is built from  $\alpha$ -glucose we call the linkages  $\alpha$  linkages; cellulose has  $\beta$  linkages.

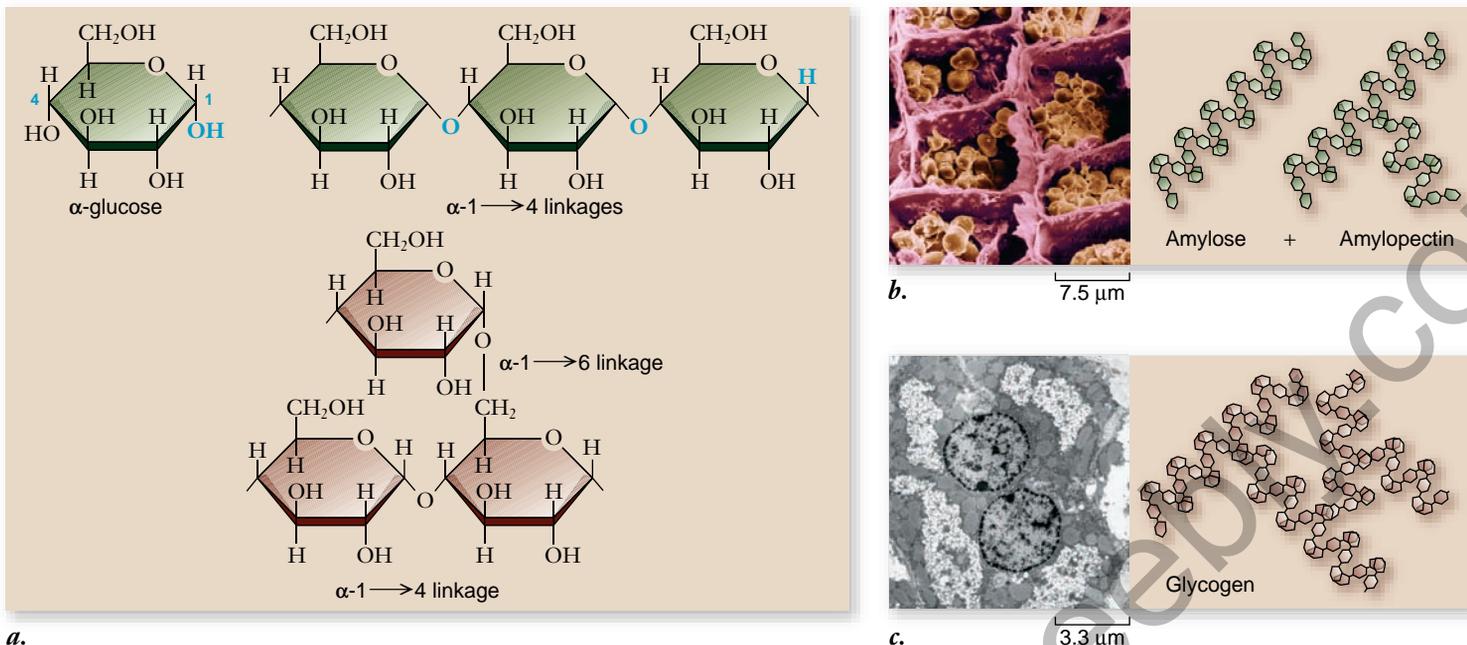
### Starches and Glycogen

Organisms store the metabolic energy contained in monosaccharides by converting them into disaccharides, such as *maltose* (figure 3.9*b*). These are then linked together into the insoluble polysaccharides called *starches*. These polysaccharides differ mainly in how the polymers branch.

The starch with the simplest structure is *amylose*. It is composed of many hundreds of  $\alpha$ -glucose molecules linked together in long, unbranched chains. Each linkage occurs between the carbon 1



**Figure 3.9 How disaccharides form.** Some disaccharides are used to transport glucose from one part of an organism's body to another; one example is sucrose (*a*), which is found in sugarcane. Other disaccharides, such as maltose (*b*), are used in grain for storage.



**Figure 3.10 Polymers of glucose: Starch and glycogen.** *a.* Starch chains consist of polymers of  $\alpha$ -glucose subunits joined by  $\alpha$ -(1 $\rightarrow$ 4) glycosidic linkages. These chains can be branched by forming similar  $\alpha$ -(1 $\rightarrow$ 6) glycosidic bonds. These storage polymers then differ primarily in their degree of branching. *b.* Starch is found in plants and is composed of amylose and amylopectin, which are unbranched and branched, respectively. The branched form is insoluble and forms starch granules in plant cells. *c.* Glycogen is found in animal cells and is highly branched and also insoluble, forming glycogen granules.

(C-1) of one glucose molecule and the C-4 of another, making them  $\alpha$ -(1 $\rightarrow$ 4) linkages (figure 3.10*a*). The long chains of amylose tend to coil up in water, a property that renders amylose insoluble. Potato starch is about 20% amylose (figure 3.10*b*).

Most plant starch, including the remaining 80% of potato starch, is a somewhat more complicated variant of amylose called *amylopectin*. Pectins are branched polysaccharides with the branches occurring due to bonds between the C-1 of one molecule and the C-6 of another [ $\alpha$ -(1 $\rightarrow$ 6) linkages]. These short amylose branches consist of 20 to 30 glucose subunits (figure 3.10*b*).

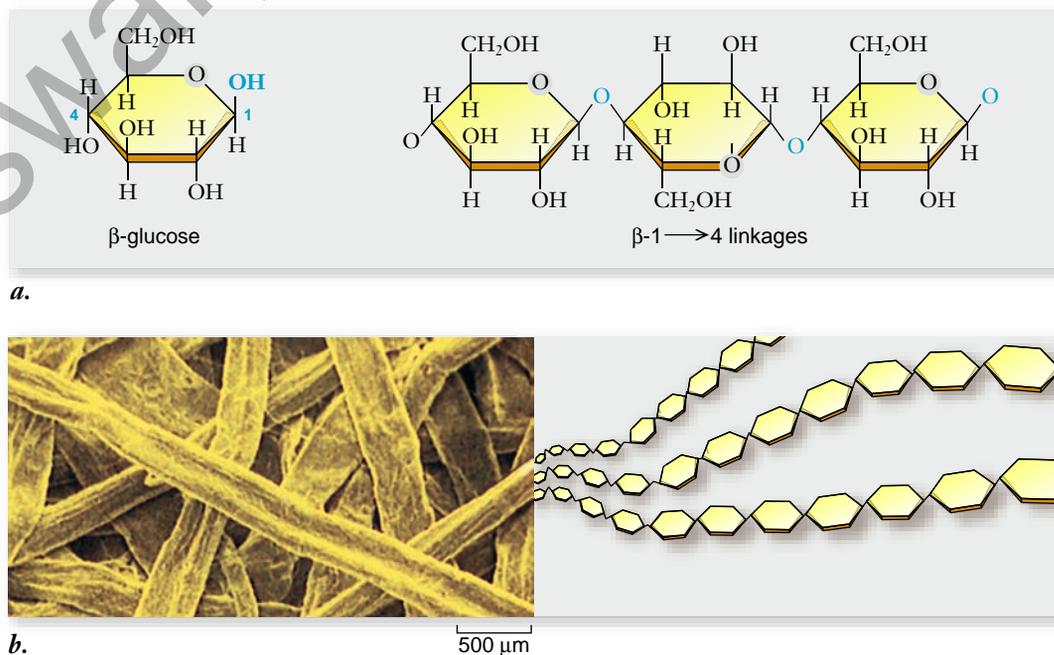
The comparable molecule to starch in animals is **glycogen**. Like amylopectin, glycogen is an insoluble polysaccharide containing branched amylose chains. Glycogen has a much longer average chain length and more branches than plant starch (figure 3.10*c*).

### Cellulose

Although some chains of sugars store energy, others serve as structural material for cells. For two glucose molecules to link together, the glucose subunits must be of the same form. *Cellulose* is a polymer of  $\beta$ -glucose (figure 3.11). The bonds between

### Figure 3.11 Polymers of glucose: Cellulose.

Starch chains consist of  $\alpha$ -glucose subunits, and cellulose chains consist of  $\beta$ -glucose subunits. *a.* Thus the bonds between adjacent glucose molecules in cellulose are  $\beta$ -(1 $\rightarrow$ 4) glycosidic linkages. *b.* Cellulose is unbranched and forms long fibers. Cellulose fibers can be very strong and are quite resistant to metabolic breakdown, which is one reason wood is such a good building material.



## 3.3 Nucleic Acids: Information Molecules

adjacent glucose molecules still exist between the C-1 of the first glucose and the C-4 of the next glucose, but these are  $\beta$ -(1 $\rightarrow$ 4) linkages.

The properties of a chain of glucose molecules consisting of all  $\beta$ -glucose are very different from those of starch. These long, unbranched  $\beta$ -linked chains make tough fibers. Cellulose is the chief component of plant cell walls (see figure 3.11*b*). It is chemically similar to amylose, with one important difference: The starch-hydrolyzing enzymes that occur in most organisms cannot break the bond between two  $\beta$ -glucose units because they only recognize  $\alpha$  linkages.

Because cellulose cannot be broken down readily by most creatures, it works well as a biological structural material. But some animals, such as cows, are able to break down cellulose by means of symbiotic bacteria and protists in their digestive tracts. These organisms provide the necessary enzymes for cleaving the  $\beta$ -(1 $\rightarrow$ 4) linkages, thus enabling access to a rich source of energy.

### Chitin

**Chitin**, the structural material found in arthropods and many fungi, is a polymer of *N*-acetylglucosamine, a substituted version of glucose. When cross-linked by proteins, it forms a tough, resistant surface material that serves as the hard exoskeleton of insects and crustaceans (figure 3.12; see chapter 34). Few organisms are able to digest chitin, but most possess a chitinase enzyme, probably to protect against fungi.

### Learning Outcomes Review 3.2

Monosaccharides have three to six or more carbon atoms typically arranged in a ring form. Disaccharides consist of two linked monosaccharides; polysaccharides are long chains of monosaccharides. Structural differences between sugar isomers can lead to functional differences. Starches are branched polymers of  $\alpha$ -glucose used for energy storage. Cellulose in plants consists of unbranched chains of  $\beta$ -glucose that are not easily digested.

- How do the structures of starch, glycogen, and cellulose affect their function?



**Figure 3.12 Chitin.** Chitin is the principal structural element in the external skeletons of many invertebrates, such as this lobster.

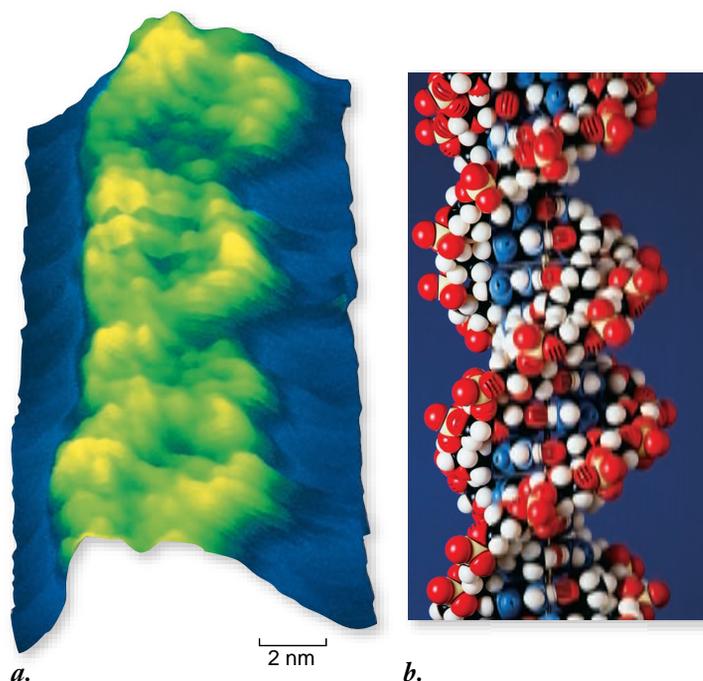
### Learning Outcomes

1. Describe the structure of nucleotides.
2. Compare and contrast the structures of DNA and RNA.
3. Explain the functions of DNA and RNA.
4. Recognize other nucleotides involved in energy metabolism.

The biochemical activity of a cell depends on production of a large number of proteins, each with a specific sequence. The information necessary to produce the correct proteins is passed through generations of organisms, even though the protein molecules themselves are not.

Nucleic acids carry information inside cells, just as disks contain the information in a computer or road maps display information needed by travelers. Two main varieties of nucleic acids are **deoxyribonucleic acid (DNA)** (figure 3.13) and **ribonucleic acid (RNA)**.

DNA encodes the genetic information used to assemble proteins (as discussed in detail in chapter 15) similar to the way the letters on this page encode information. Unique among macromolecules, nucleic acids are able to serve as templates to produce precise copies of themselves. This characteristic allows genetic information to be preserved during cell division and



**Figure 3.13 Images of DNA.** *a.* A scanning-tunneling micrograph of DNA (false color; 2,000,000 $\times$ ) showing approximately three turns of the DNA double helix. *b.* A space-filling model for comparison to the image of actual DNA in (*a*).

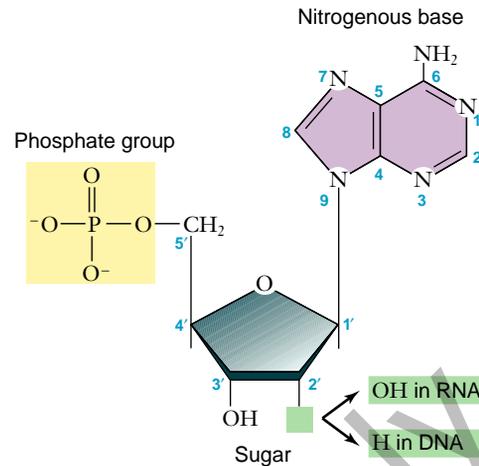
during the reproduction of organisms. DNA, found primarily in the nuclear region of cells, contains the genetic information necessary to build specific organisms.

Cells use a type of RNA called messenger RNA (mRNA) to direct the synthesis of proteins. mRNA consists of transcribed single-stranded copies of portions of the DNA. These transcripts serve as blueprints specifying the amino acid sequences of proteins. This process will be described in detail in chapter 15.

## Nucleic acids are nucleotide polymers

Nucleic acids are long polymers of repeating subunits called **nucleotides**. Each nucleotide consists of three components: a pentose, or five-carbon sugar (ribose in RNA and deoxyribose in DNA); a phosphate ( $-\text{PO}_4$ ) group; and an organic nitrogenous (nitrogen-containing) base (figure 3.14). When a nucleic acid polymer forms, the phosphate group of one nucleotide binds to the hydroxyl group from the pentose sugar of another, releasing water and forming a *phosphodiester bond* by a dehydration reaction. A **nucleic acid**, then, is simply a chain of five-carbon sugars linked together by phosphodiester bonds with a nitrogenous base protruding from each sugar (see figure 3.15a). These chains of nucleotides, *polynucleotides*, have different ends: a phosphate on one end and an  $-\text{OH}$  from a sugar on the other end. We conventionally refer to these ends as 5' ("five-prime,"  $-\text{PO}_4$ ) and 3' ("three-prime,"  $-\text{OH}$ ) taken from the carbon numbering of the sugar (figure 3.15a).

Two types of nitrogenous bases occur in nucleotides (3.15b). The first type, *purines*, are large, double-ring molecules found in both DNA and RNA; the two types of purines are adenine (A) and guanine (G). The second type, *pyrimidines*, are smaller, single-ring molecules; they include cytosine (C, in both

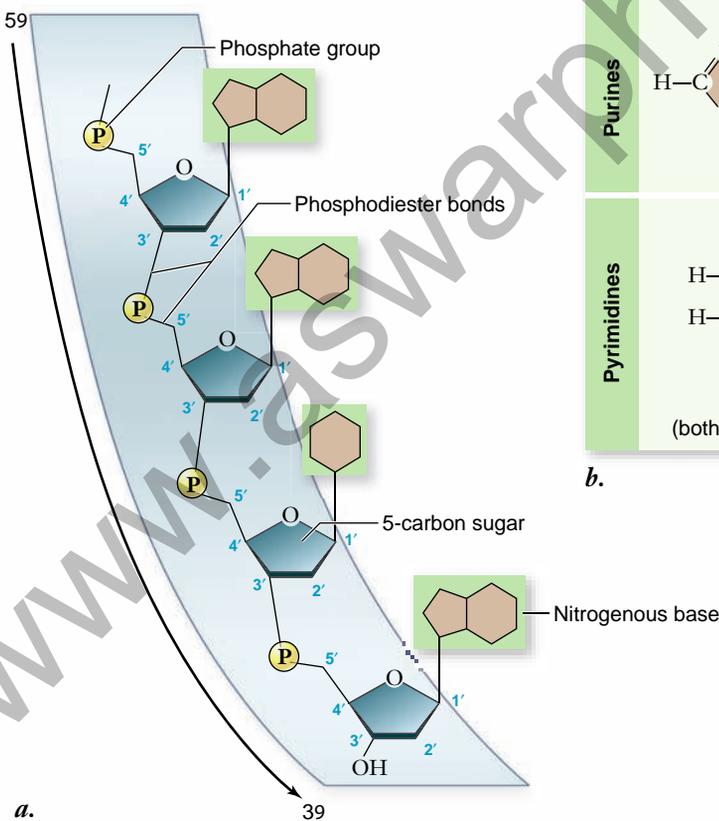


**Figure 3.14 Structure of a nucleotide.** The nucleotide subunits of DNA and RNA are made up of three elements: a five-carbon sugar (ribose or deoxyribose), an organic nitrogenous base (adenine is shown here), and a phosphate group. Notice that all the numbers on the sugar are given as "primes" (1', 2', etc.) to distinguish them from the numbering on the rings of the bases.

DNA and RNA), thymine (T, in DNA only), and uracil (U, in RNA only).

## DNA carries the genetic code

Organisms use sequences of nucleotides in DNA to encode the information specifying the amino acid sequences of their proteins. This method of encoding information is very similar to the way in which sequences of letters encode information in a sentence.



**Figure 3.15 The structure of a nucleic acid and the organic nitrogenous bases.** *a.* In a nucleic acid, nucleotides are linked to one another via phosphodiester bonds formed between the phosphate of one nucleotide and the sugar of the next nucleotide. We call this the sugar-phosphate backbone, and the organic bases protrude from this chain. The backbone also has different ends: a 5' phosphate end and a 3' hydroxyl end (the numbers come from the numbers in the sugars). *b.* The organic nitrogenous bases can be either purines or pyrimidines. The base thymine is found in DNA. The base uracil is found in RNA.