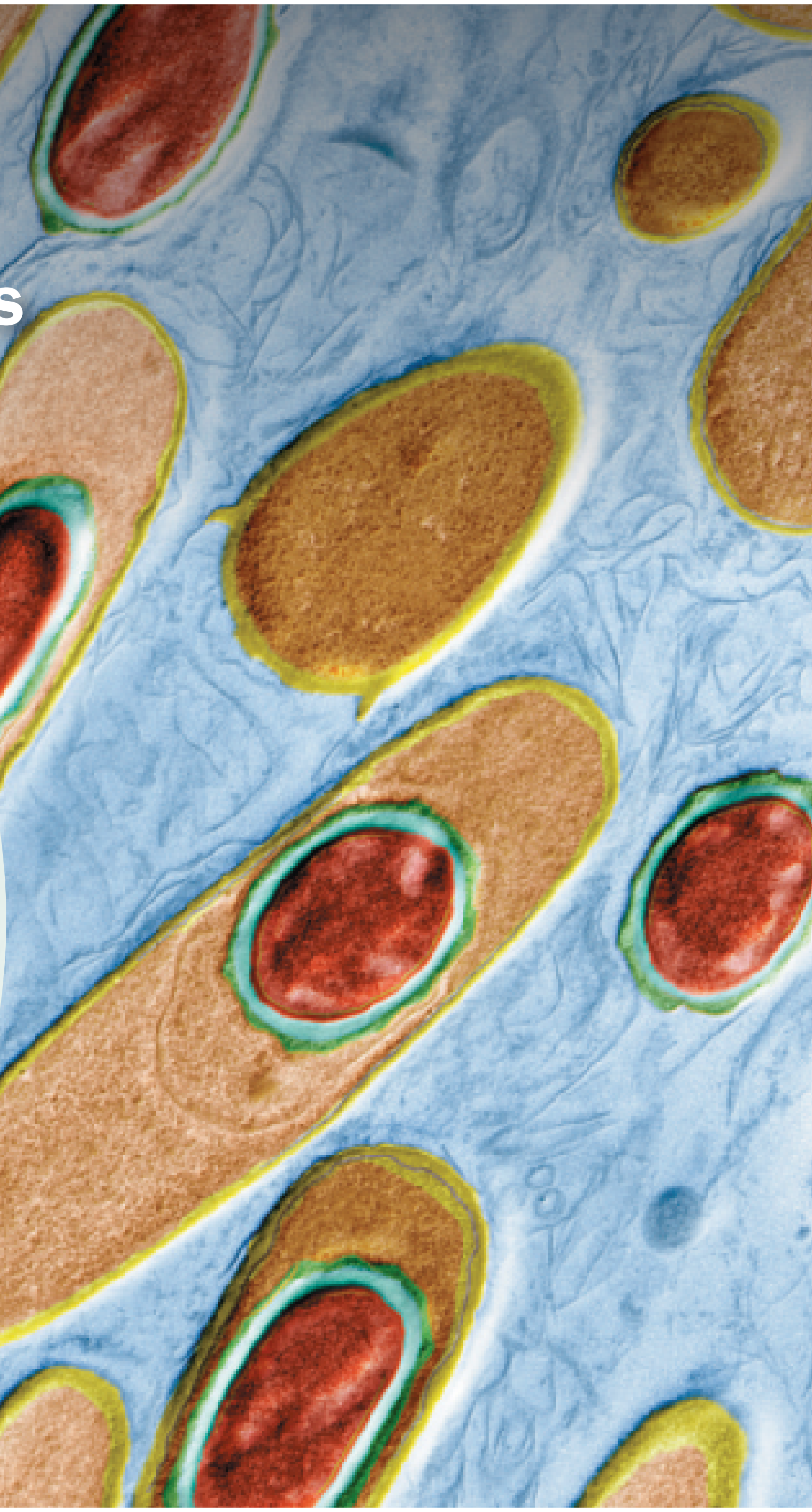
2 Chemical Principles

Like all organisms, microorganisms use nutrients to make chemical building blocks for growth and other functions essential to life. For most microorganisms, synthesizing these building blocks requires them to break down nutrient substances and use the energy released to assemble the resulting molecular fragments into new substances.

We see evidence of these microbial chemical reactions in the world routinely, from a fallen tree rotting on the forest floor to milk going sour in the refrigerator. Although most people give little thought to the causes of such things, the chemistry of microbes is one of the most important concerns of microbiologists. Knowledge of chemistry is essential to understanding what roles microorganisms play in nature, how they cause disease, how methods for diagnosing disease are developed, how the body's defenses combat infection, and how antibiotics and vaccines are produced to combat the harmful effects of microbes.

The *Bacillus anthracis* (bah-SIL-lus an-THRĀ-sis) bacteria in the photograph make a capsule that is not readily digested by animal cells. As discussed in the Clinical Case, these bacteria can grow in mammals by avoiding host defenses. Researchers are investigating ways to identify unique chemicals made by *B. anthracis* and other potential biological weapons in order to detect bioterrorism. To understand the changes that occur in microorganisms and the changes microbes make in the world around us, we need to know how molecules are formed and how they interact.



▶ ***Bacillus anthracis*** **bacteria produce heat-resistant endospores (red).**

In the Clinic

As the health advisory nurse at a health service company, you receive a call from a man who is concerned that his blood sugar levels have not decreased, even though he has switched to using organic sugar. **How would you respond to the man?**

Hint: Read about important biological molecules later in this chapter on pages 31–47.

Play In the Clinic Video @MasteringMicrobiology

The Structure of Atoms

LEARNING OBJECTIVE

2-1 Describe the structure of an atom and its relation to the physical properties of elements.

All matter—whether air, rock, or a living organism—is made up of small units called atoms. An **atom** is the smallest component of a substance, and it cannot be subdivided into smaller substances without losing its properties. Atoms combine to form **molecules**. Living cells are made up of molecules, some of which are very complex. The science of the interaction between atoms and molecules is called **chemistry**.

Atoms are the smallest units of matter that enter into chemical reactions. Every atom has a centrally located **nucleus** and negatively (−) charged particles called **electrons** that move around the nucleus in regions called *electron shells* (Figure 2.1). The nucleus is made up of positively (+) charged particles called **protons** and uncharged (neutral) particles called **neutrons**. The nucleus, therefore, bears a net positive charge. All atoms contain an equal number of electrons and protons. Because the total positive charge of the nucleus equals the total negative charge of the electrons, each atom is electrically neutral.

The nuclei of most atoms are stable—that is, they do not change spontaneously—and nuclei do not participate in chemical reactions. The number of protons in an atomic nucleus ranges from one (in a hydrogen atom) to more than 100 (in the largest atoms known). Atoms are often listed by their **atomic number**, the number of protons in the nucleus. Protons and neutrons are approximately the same weight, which is about 1840 times that of an electron, and the total number of protons and neutrons in an atom is its approximate **atomic mass**.

Chemical Elements

All atoms with the same number of protons behave the same way chemically and are classified as the same **chemical element**. Each element has its own name and a one- or two-letter symbol, usually derived from the English or Latin name for the element. For example, the symbol for the element hydrogen is H, and the symbol for carbon is C. The symbol for sodium is Na—the first two letters of its Latin name, *natrium*—distinguish it from nitrogen, N, and from sulfur, S. There are 92 naturally occurring elements. However, only about 26 elements are commonly found in living things. Table 2.1 lists some of the chemical elements found in living organisms.

Most elements have several **isotopes**—atoms with different numbers of neutrons in their nuclei. All isotopes of an element have the same number of protons in their nuclei, but their atomic masses differ because of the difference in the number of neutrons. For example, in a natural sample of oxygen, all the atoms contain eight protons. However, 99.76% of the atoms have eight neutrons, 0.04% contain nine neutrons, and the remaining 0.2% contain ten neutrons. Therefore, the three isotopes composing a natural sample of oxygen have atomic masses of 16, 17, and 18, although all will have the atomic number 8. Atomic numbers are written as a subscript to the left of an element's chemical symbol. Atomic masses are written as a superscript above the atomic number. Thus, natural oxygen isotopes are represented as ${}^{16}_{8}O$, ${}^{17}_{8}O$, and ${}^{18}_{8}O$. Isotopes of certain elements are extremely useful in biological research, medical diagnosis, the treatment of some disorders, and some forms of sterilization.

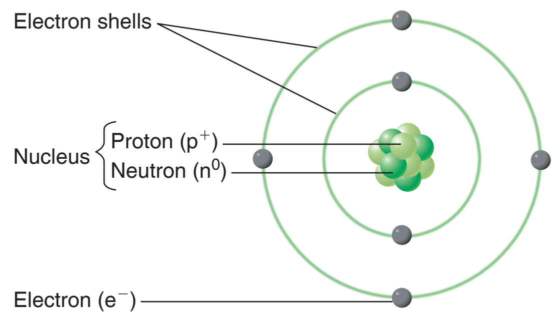


Figure 2.1 The structure of an atom. In this simplified diagram of a carbon atom, note the central location of the nucleus. The nucleus contains six neutrons and six protons, although not all the neutrons are visible in this view. The six electrons move about the nucleus in regions called electron shells, shown here as circles.

Q What is the atomic number of this atom?

CLINICAL CASE Drumming Up Dust

Jonathan, a 52-year-old drummer, is doing his best to ignore the cold sweat that is breaking out all over his body. He and his bandmates are performing in a local Philadelphia nightclub, and they are just about finished with the second set of the evening. Jonathan hasn't been feeling well for a while, actually; he has been feeling weak and short of breath for the last 3 days or so. Jonathan makes it to the end of the song, but the noise from the clapping and cheering audience seems to come from far away. He stands up to bow and collapses. Jonathan is admitted to a local emergency department with a mild fever and severe shaking. He is able to tell the admitting nurse that he also has had a dry cough for the last few days. The attending physician orders a chest X-ray exam and sputum culture. Jonathan is diagnosed with bilateral pneumonia caused by *Bacillus anthracis*. The attending physician is astonished by this diagnosis.

How did Jonathan become infected by *B. anthracis*? Read on to find out.

25 42 44 46

TABLE 2.1 The Elements of Life*

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

*Hydrogen, carbon, nitrogen, and oxygen are the most abundant chemical elements in living organisms.

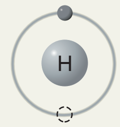
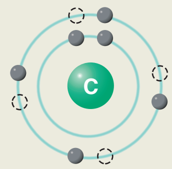
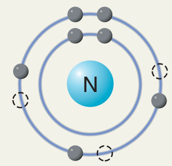
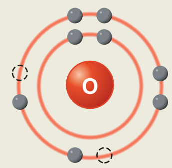
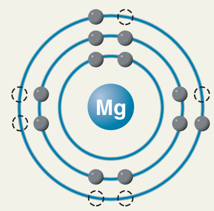
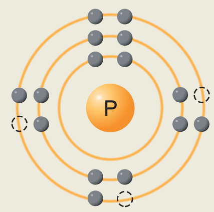
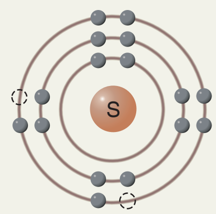
Electronic Configurations

In an atom, electrons are arranged in **electron shells**, which are regions corresponding to different **energy levels**. The arrangement is called an **electronic configuration**. Shells are layered outward from the nucleus, and each shell can hold a characteristic maximum number of electrons—two electrons in the innermost shell (lowest energy level), eight electrons in the second shell, and eight electrons in the third shell, if it is the atom's outermost (valence) shell. The fourth, fifth, and sixth electron shells can each accommodate 18 electrons, although there are some exceptions to this generalization. Table 2.2 shows the electronic configurations for atoms of some elements found in living organisms.

The number of electrons in the outermost shell determines an atom's tendency to react with other atoms. An atom can give up, accept, or share electrons with other atoms to fill the outermost shell. When its outer shell is filled, the atom is chemically stable, or inert: it does not tend to react with other atoms. Helium (atomic number 2) and neon (atomic number 10) are examples of atoms of inert gases whose outer shells are filled.

When an atom's outer electron shell is only partially filled, the atom is chemically unstable. These unstable atoms react with other atoms, depending, in part, on the degree to which the outer energy levels are filled. Notice the number of electrons in the outer energy levels of the atoms in Table 2.2. We will see later how the number correlates with the chemical reactivity of the elements.

TABLE 2.2 Electronic Configurations for the Atoms of Some Elements Found in Living Organisms

Element	Diagram	Number of Valence (Outermost) Shell Electrons	Number of Unfilled Spaces	Maximum Number of Bonds Formed
Hydrogen		1	1	1
Carbon		4	4	4
Nitrogen		5	3	5
Oxygen		6	2	2
Magnesium		2	6	2
Phosphorus		5	3	5
Sulfur		6	2	6

CHECK YOUR UNDERSTANDING

2-1 How does $^{14}_{6}C$ differ from $^{12}_{6}C$? What is the atomic number of each carbon atom? The atomic mass?

How Atoms Form Molecules: Chemical Bonds

LEARNING OBJECTIVE

2-2 Define *ionic bond*, *covalent bond*, *hydrogen bond*, *molecular weight*, and *mole*.

When the outermost energy level of an atom is not completely filled by electrons, you can think of it as having either unfilled spaces or extra electrons in that energy level, depending on whether it is easier for the atom to gain or lose electrons. For example, an atom of oxygen, with two electrons in the first energy level and six in the second, has two unfilled spaces in the second electron shell; an atom of magnesium has two extra electrons in its outermost shell. The most chemically stable configuration for any atom is to have its outermost shell filled. Therefore, for these two atoms to attain that state, oxygen must gain two electrons, and magnesium must lose two electrons. Because all atoms tend to combine so that the extra electrons in the outermost shell of one atom fill the spaces of the outermost shell of the other atom, oxygen and magnesium combine so that the outermost shell of each atom has the full complement of eight electrons.

The **valence**, or combining capacity, of an atom is the number of extra or missing electrons in its outermost electron shell. For example, hydrogen has a valence of 1 (one unfilled space, or one extra electron), oxygen has a valence of 2 (two unfilled spaces), carbon has a valence of 4 (four unfilled spaces, or four extra electrons), and magnesium has a valence of 2 (two extra electrons).

Basically, atoms achieve the full complement of electrons in their outermost energy shells by combining to form molecules, which are made up of atoms of one or more elements. A molecule that contains at least two different kinds of atoms, such as H_2O (the water molecule), is called a **compound**. In H_2O, the subscript 2 indicates that there are two atoms of hydrogen; the absence of a subscript indicates that there is only one atom of oxygen. Molecules hold together because the valence electrons of the combining atoms form attractive forces, called **chemical bonds**, between the atomic nuclei. Therefore, valence may also be viewed as the bonding capacity of an element.

In general, atoms form bonds in one of two ways: by either gaining or losing electrons from their outer electron shell, or by sharing outer electrons. When atoms have gained or lost outer electrons, the chemical bond is called an *ionic bond*. When outer electrons are shared, the bond is called a *covalent bond*. Although we will discuss ionic and covalent bonds separately, the kinds of bonds actually found in molecules do not belong entirely to either category. Instead, bonds range from the highly ionic to the highly covalent.

Ionic Bonds

Atoms are electrically neutral when the number of positive charges (protons) equals the number of negative charges (electrons). But when an isolated atom gains or loses electrons, this balance is upset. If the atom gains electrons, it acquires an overall negative charge; if the atom loses electrons, it acquires an overall positive charge. Such a negatively or positively charged atom (or group of atoms) is called an **ion**.

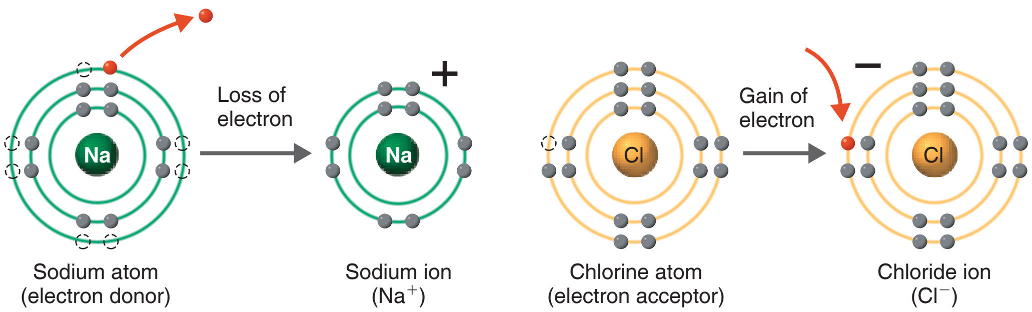
Consider the following examples. Sodium (Na) has 11 protons and 11 electrons, with one electron in its outer electron shell. Sodium tends to lose the single outer electron; it is an *electron donor* (Figure 2.2a). When sodium donates an electron to another atom, it is left with 11 protons and only 10 electrons and so has an overall charge of +1. This positively charged sodium atom is called a sodium ion and is written as Na^+. Chlorine (Cl) has a total of 17 electrons, seven of them in the outer electron shell. Because this outer shell can hold eight electrons, chlorine tends to pick up an electron that has been lost by another atom; it is an *electron acceptor* (see Figure 2.2a). By accepting an electron, chlorine totals 18 electrons. However, it still has only 17 protons in its nucleus. The chloride ion therefore has a charge of −1 and is written as Cl^-.

The opposite charges of the sodium ion (Na^+) and chloride ion (Cl^-) attract each other. The attraction, an ionic bond, holds the two atoms together, and a molecule is formed (Figure 2.2b). The formation of this molecule, called sodium chloride (NaCl) or table salt, is a common example of ionic bonding. Thus, an **ionic bond** is an attraction between ions of opposite charge that holds them together to form a stable molecule. Put another way, an ionic bond is an attraction between atoms in which one atom loses electrons and another atom gains electrons. Strong ionic bonds, such as those that hold Na^+ and Cl^- together in salt crystals, have limited importance in living cells. But the weaker ionic bonds formed in aqueous (water) solutions are important in biochemical reactions in microbes and other organisms. For example, weaker ionic bonds assume a role in certain antigen–antibody reactions—that is, reactions in which molecules produced by the immune system (antibodies) combine with foreign substances (antigens) to combat infection.

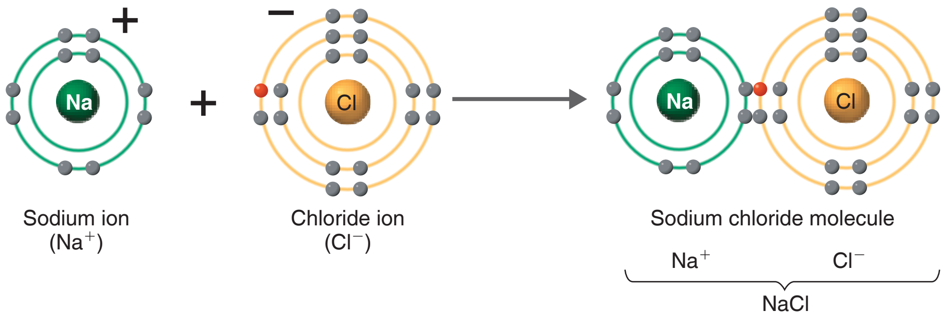
In general, an atom whose outer electron shell is less than half-filled will lose electrons and form positively charged ions, called **cations**. Examples of cations are the potassium ion (K^+), calcium ion (Ca^{2+}), and sodium ion (Na^+). When an atom's outer electron shell is more than half-filled, the atom will gain electrons and form negatively charged ions, called **anions**. Examples are the iodide ion (I^-), chloride ion (Cl^-), and sulfide ion (S^{2-}).

Covalent Bonds

A **covalent bond** is a chemical bond formed by two atoms sharing one or more pairs of electrons. Covalent bonds are stronger



(a) A sodium atom (Na) loses one electron to an electron acceptor and forms a sodium ion (Na^+). A chlorine atom (Cl) accepts one electron from an electron donor to become a chloride ion (Cl^-).



(b) The sodium and chloride ions are attracted because of their opposite charges and are held together by an ionic bond to form a molecule of sodium chloride.

Figure 2.2 Ionic bond formation.

Q What is an ionic bond?

and far more common in organisms than are true ionic bonds. In the hydrogen molecule, H_2, two hydrogen atoms share a pair of electrons. Each hydrogen atom has its own electron plus one electron from the other atom (Figure 2.3a). The shared pair of electrons actually orbits the nuclei of both atoms. Therefore, the outer electron shells of both atoms are filled. Atoms that share only one pair of electrons form a *single covalent bond*. For simplicity, a single covalent bond is expressed as a single line between the atoms (H—H). Atoms that share two pairs of electrons form a *double covalent bond*, expressed as two single lines (═). A *triple covalent bond*, expressed as three single lines (≡), occurs when atoms share three pairs of electrons.

The principles of covalent bonding that apply to atoms of the same element also apply to atoms of different elements. Methane (CH_4) is an example of covalent bonding between atoms of different elements (Figure 2.3b). The outer electron shell of the carbon atom can hold eight electrons but has only four; each hydrogen atom can hold two electrons but has only one. Consequently, in the methane molecule the carbon atom gains four hydrogen electrons to complete its outer shell, and each hydrogen atom completes its pair by sharing one electron from the carbon atom. Each outer electron of the carbon atom orbits both the carbon nucleus and a hydrogen nucleus. Each hydrogen electron orbits both its own nucleus and the carbon nucleus.

Elements such as hydrogen and carbon, whose outer electron shells are half-filled, form covalent bonds quite easily. In fact, in living organisms, carbon almost always forms covalent bonds; it almost never becomes an ion. *Remember*: Covalent bonds are formed by the *sharing* of electrons between atoms. Ionic bonds are formed by *attraction* between atoms that have lost or gained electrons and are therefore positively or negatively charged.

Hydrogen Bonds

Another chemical bond of special importance to all organisms is the **hydrogen bond**, in which a hydrogen atom that is covalently bonded to one oxygen or nitrogen atom is attracted to another oxygen or nitrogen atom. Such bonds are weak and do not bind atoms into molecules. However, they do serve as bridges between different molecules or between various portions of the same molecule.

When hydrogen combines with atoms of oxygen or nitrogen, the relatively large nucleus of these larger oxygen or nitrogen atoms has more protons and attracts the hydrogen electron more strongly than does the small hydrogen nucleus. Thus, in a molecule of water (H_2O), all the electrons tend to be closer to the oxygen nucleus than to the hydrogen nuclei. As a result, the oxygen portion of the molecule has a slightly negative charge, and the hydrogen portion of the molecule has a slightly positive charge (Figure 2.4a). When the positively charged end of one

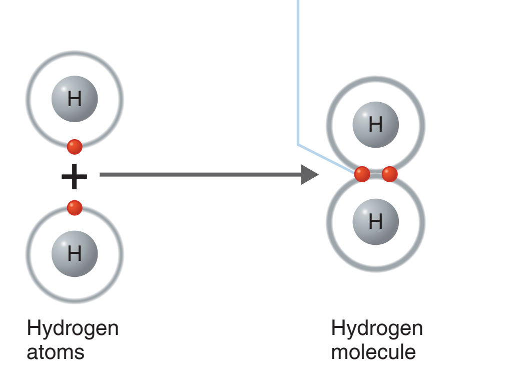
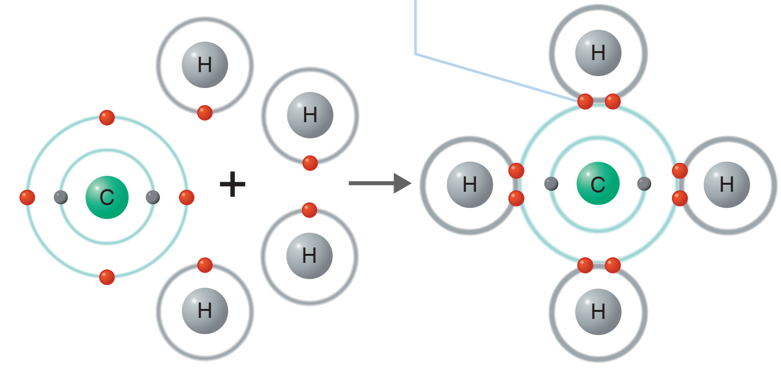
	(a) Hydrogen	(b) Methane
Molecular formula	H_2	CH_4
Structural formula	H—H; A single covalent bond forms between two hydrogen atoms to form a hydrogen molecule.	H–C(–H)(–H)–H; Single covalent bonds between four hydrogen atoms and a carbon atom form a methane molecule.
Atomic diagram		

Figure 2.3 Covalent bond formation. The molecular formula shows the number and types of atoms in a molecule. In structural formulas, each covalent bond is written as a straight line between the symbols for two atoms. In molecular formulas, the number of atoms in each molecule is noted by subscripts.

Q What is a covalent bond?

molecule is attracted to the negatively charged end of another molecule, a hydrogen bond is formed (Figure 2.4b). This attraction can also occur between hydrogen and other atoms of the same molecule, especially in large molecules, but oxygen and nitrogen are the elements most frequently involved in hydrogen bonding.

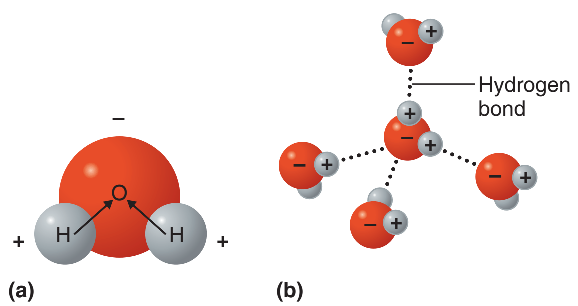


Figure 2.4 Hydrogen bond formation in water. (a) In a water molecule, the electrons of the hydrogen atoms are strongly attracted to the oxygen atom. Therefore, the part of the water molecule containing the oxygen atom has a slightly negative charge, and the part containing hydrogen atoms has a slightly positive charge. **(b)** In a hydrogen bond between water molecules, the hydrogen of one water molecule is attracted to the oxygen of another water molecule. Many water molecules may be attracted to each other by hydrogen bonds (black dots).

Q Which chemical elements are usually involved in hydrogen bonding?

Hydrogen bonds are considerably weaker than either ionic or covalent bonds; they have only about 5% of the strength of covalent bonds. Consequently, hydrogen bonds are formed and broken relatively easily. This property accounts for the temporary bonding that occurs between certain atoms of large and complex molecules, such as proteins and nucleic acids. Even though hydrogen bonds are relatively weak, large molecules containing several hundred of these bonds have considerable strength and stability. A summary of ionic, covalent, and hydrogen bonds is shown in Table 2.3.

Molecular Mass and Moles

You have seen that bond formation usually results in the creation of molecules. Molecules are often discussed in terms of units of measure called molecular mass and moles. The **molecular mass** of a molecule is the sum of the atomic masses of all its atoms. To relate the molecular level to the laboratory level, we use a unit called the mole. One **mole** of a substance is its molecular mass expressed in grams. The unit of molecular mass is a **dalton (da)**. For example, 1 mole of water weighs 18 grams because the molecular mass of H_2O is 18 da, or $[(2 \times 1) + 16]$.

CHECK YOUR UNDERSTANDING

✓ **2-2** Differentiate an ionic bond from a covalent bond.

TABLE 2.3 Comparison among Ionic, Covalent, and Hydrogen Bonds

Type of Bond	Definition and Importance
Ionic	An *attraction* between ions of opposite charge that holds them together to form a stable molecule. Weaker ionic bonds are important in biochemical reactions such as antigen–antibody reactions.
Covalent	A bond formed by two atoms that *share* one or more pairs of electrons. Covalent bonds are the most common type of chemical bond in organisms and are responsible for holding together the atoms of most molecules in organisms.
Hydrogen	A relatively weak bond in which a hydrogen atom that is covalently bonded to one oxygen or nitrogen atom is attracted to another oxygen or nitrogen atom. Hydrogen bonds do not bind atoms into molecules, but serve as *bridges between different molecules* or different portions of the same molecule, for example, within proteins and nucleic acids.

Chemical Reactions

LEARNING OBJECTIVE

2-3 Diagram three basic types of chemical reactions.

As we said earlier, **chemical reactions** involve the making or breaking of bonds between atoms. After a chemical reaction, the total number of atoms remains the same, but there are new molecules with new properties because the atoms have been rearranged.

Energy in Chemical Reactions

All chemical bonds require energy to form or break. It is important to note that initially, *activation energy* is needed to break a bond (see page 111). In the chemical reactions of metabolism, energy is released when new bonds are formed after the original bonds break; this is the energy cells use to do work. A chemical reaction that absorbs more energy than it releases is called an **endergonic reaction** (*endo* = within), meaning that energy is directed inward. A chemical reaction that releases more energy than it absorbs is called an **exergonic reaction** (*exo* = out), meaning that energy is directed outward.

In this section we will look at three basic types of chemical reactions common to all living cells. By becoming familiar with these reactions, you will be able to understand the specific chemical reactions we will discuss later (particularly in Chapter 5).

Synthesis Reactions

When two or more atoms, ions, or molecules combine to form new and larger molecules, the reaction is called a **synthesis reaction**. To synthesize means to put together, and a synthesis reaction *forms new bonds*. Synthesis reactions can be expressed in the following way:

$$\underbrace{\underset{\text{Atom, ion, or molecule A}}{A} + \underset{\text{Atom, ion, or molecule B}}{B}}_{\textit{reactants}} \xrightarrow{\textit{combine to form}} \underbrace{\underset{\text{New molecule AB}}{AB}}_{\textit{product}}$$

The combining substances, A and B, are called the *reactants*; the substance formed by the combination, AB, is the *product*. The *arrow* indicates the direction in which the reaction proceeds.

Pathways of synthesis reactions in living organisms are collectively called anabolic reactions, or simply **anabolism** (an-AB-ō-liz-um). The combining of sugar molecules to form starch and of amino acids to form proteins are two examples of anabolism.

Decomposition Reactions

The reverse of a synthesis reaction is a **decomposition reaction**. To decompose means to break down into smaller parts, and in a decomposition reaction *bonds are broken*. Typically, decomposition reactions split large molecules into smaller molecules, ions, or atoms. A decomposition reaction occurs in the following way:

$$\underbrace{\underset{\text{Molecule AB}}{AB}}_{\textit{reactant}} \xrightarrow{\textit{breaks down into}} \underbrace{\underset{\text{Atom, ion, or molecule A}}{A} + \underset{\text{Atom, ion, or molecule B}}{B}}_{\textit{products}}$$

Decomposition reactions that occur in living organisms are collectively called catabolic reactions, or simply **catabolism** (ka-TAB-ō-liz-um). An example of catabolism is the breakdown of sucrose (table sugar) into simpler sugars, glucose and fructose, during digestion.

Exchange Reactions

All chemical reactions are based on synthesis and decomposition. Many reactions, such as **exchange reactions**, are actually part synthesis and part decomposition. An exchange reaction works in the following way:

$$\underbrace{AB + CD}_{\textit{reactants}} \xrightarrow{\textit{recombine to form}} \underbrace{AD + BC}_{\textit{products}}$$

First, the bonds between A and B and between C and D are broken in a decomposition process. New bonds are then formed between A and D and between B and C in a synthesis process. For example, an exchange reaction occurs when sodium hydroxide (NaOH) and hydrochloric acid (HCl) react to form table salt (NaCl) and water (H_2O), as follows:

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

The Reversibility of Chemical Reactions

All chemical reactions are, in theory, reversible; that is, they can occur in either direction. In practice, however, some reactions do this more easily than others. A chemical reaction that is readily reversible (when the end product can revert to the original molecules) is termed a **reversible reaction** and is indicated by two arrows, as shown here:

$$\underset{A + B}{\textit{reactants}} \quad \underset{\text{break down into}}{\overset{\text{combine to form}}{\rightleftharpoons}} \quad \underset{AB}{\textit{product}}$$

Some reversible reactions occur because neither the reactants nor the end products are very stable. Other reactions reverse only under special conditions:

$$\underset{A + B}{\textit{reactants}} \quad \underset{\text{water}}{\overset{\text{heat}}{\rightleftharpoons}} \quad \underset{AB}{\textit{product}}$$

Whatever is written above or below the arrows indicates the special condition under which the reaction in that direction occurs. In this case, A and B react to produce AB only when heat is applied, and AB breaks down into A and B only in the presence of water. See Figure 2.8 on page 35 for another example.

In Chapter 5 we will examine the various factors that affect chemical reactions.

CHECK YOUR UNDERSTANDING

✓ **2-3** The chemical reaction below is used to remove chlorine from water. What type of reaction is it?

$$HClO + Na_2SO_3 \longrightarrow Na_2SO_4 + HCl$$

Important Biological Molecules

Biologists and chemists divide compounds into two principal classes: inorganic and organic. **Inorganic compounds** are defined as molecules, usually small and structurally simple, which typically lack carbon and in which ionic bonds may play an important role. Inorganic compounds include water, molecular oxygen (O_2), carbon dioxide, and many salts, acids, and bases.

Organic compounds always contain carbon and hydrogen and typically are structurally complex. Carbon is a unique element because it has four electrons in its outer shell and four unfilled spaces. It can combine with a variety of atoms, including other carbon atoms, to form straight or branched chains and rings. Carbon chains form the basis of many organic compounds in living cells, including sugars, amino acids, and vitamins. Organic compounds are held together mostly or entirely by covalent bonds. Some organic molecules, such as polysaccharides, proteins, and nucleic acids, are very large and usually contain thousands of atoms. Such giant molecules are called *macromolecules*. In the following section we will discuss inorganic and organic compounds that are essential for cells.

Inorganic Compounds

LEARNING OBJECTIVES

2-4 List several properties of water that are important to living systems.

2-5 Define *acid*, *base*, *salt*, and *pH*.

Water

All living organisms require a wide variety of inorganic compounds for growth, repair, maintenance, and reproduction. Water is one of the most important, as well as one of the most abundant, of these compounds, and it is particularly vital to microorganisms. Outside the cell, nutrients are dissolved in water, which facilitates their passage through cell membranes. Inside the cell, water is the medium for most chemical reactions. In fact, water is by far the most abundant component of almost all living cells. Water makes up between 65% and 75% of every cell on average. Simply stated, no organism can survive without water.

Water has structural and chemical properties that make it particularly suitable for its role in living cells. As we discussed, the total charge on the water molecule is neutral, but the oxygen region of the molecule has a slightly negative charge, and the hydrogen region has a slightly positive charge (see Figure 2.4a). Any molecule having such an unequal distribution of charges is called a **polar molecule**. The polar nature of water gives it four characteristics that make it a useful medium for living cells.

First, every water molecule is capable of forming four hydrogen bonds with nearby water molecules (see Figure 2.4b). This property results in a strong attraction between water molecules and makes water an excellent temperature buffer. Because of this strong attraction, a great deal of heat is required to separate water molecules from each other to form water vapor; thus, water has a relatively high boiling point (100°C, 212°F). Because water has such a high boiling point, it exists in the liquid state on most of the Earth's surface. Conversely, water temperatures must drop significantly in order for it to freeze. Secondly, the hydrogen bonding between water molecules affects the density of water, depending on whether it occurs as ice or a liquid. For example, the hydrogen bonds in the crystalline structure of water (ice) make ice take up more space. As a result, ice has fewer molecules than an equal volume of liquid water. This makes its crystalline structure less dense than liquid water. For this reason, ice floats and can serve as an insulating layer on the surfaces of lakes and streams that harbor living organisms.

Third, the polarity of water makes it an excellent dissolving medium, or **solvent**. Many polar substances undergo **dissociation**, or separation, into individual molecules in water—that is, they dissolve. The negative part of the water molecules

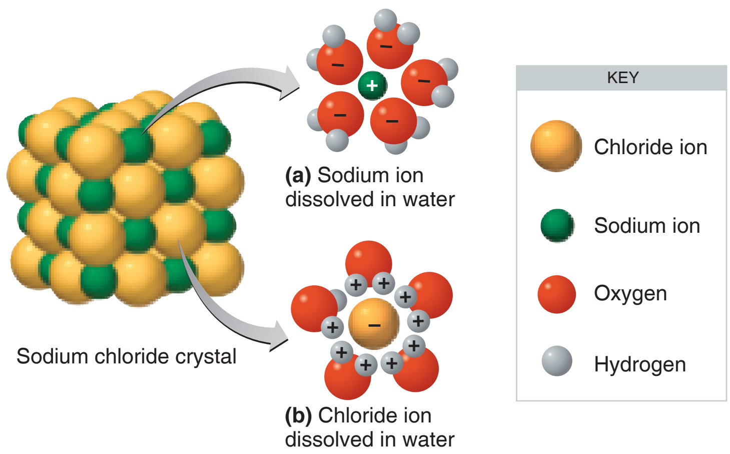


Figure 2.5 How water acts as a solvent for sodium chloride (NaCl). (a) The positively charged sodium ion (Na^+) is attracted to the negative part of the water molecule. **(b)** The negatively charged chloride ion (Cl^-) is attracted to the positive part of the water molecule. In the presence of water molecules, the bonds between the Na^+ and Cl^- are disrupted, and the NaCl dissolves in the water.

Q What happens during ionization?

is attracted to the positive part of the molecules in the **solute**, or dissolving substance, and the positive part of the water molecules is attracted to the negative part of the solute molecules. Substances (such as salts) that are composed of atoms (or groups of atoms) held together by ionic bonds tend to dissociate into separate cations and anions in water. Thus, the polarity of water allows molecules of many different substances to separate and become surrounded by water molecules (Figure 2.5).

Fourth, polarity accounts for water's characteristic role as a reactant or product in many chemical reactions. Its polarity facilitates the splitting and rejoining of hydrogen ions (H^+) and hydroxide ions (OH^-). Water is a key reactant in the digestive processes of organisms, whereby larger molecules are broken down into smaller ones. Water molecules are also involved in synthetic reactions; water is an important source of the hydrogen and oxygen that are incorporated into numerous organic compounds in living cells.

Acids, Bases, and Salts

As we saw in Figure 2.5, when inorganic salts such as sodium chloride (NaCl) are dissolved in water, they undergo **ionization** or *dissociation*; that is, they break apart into ions. Substances called acids and bases show similar behavior.

An **acid** can be defined as a substance that dissociates into one or more hydrogen ions (H^+) and one or more negative ions (anions). Thus, an acid can also be defined as a proton (H^+) donor. A **base** dissociates into one or more negatively charged hydroxide ions (OH^-) that can accept, or combine with, protons, and one or more positive ions (cations). Thus, sodium hydroxide (NaOH) is a base because it dissociates to release OH^-, which has a strong attraction for protons and is among the most important proton acceptors. A **salt** is a substance that dissociates in water into cations and anions, neither of which is H^+ or OH^-. Figure 2.6 shows common examples of each type of compound and how they dissociate in water.

Acid–Base Balance: The Concept of pH

An organism must maintain a fairly constant balance of acids and bases to remain healthy. For example, if a particular acid or base concentration is too high or too low, enzymes change in shape and no longer effectively promote chemical reactions in a cell. In the aqueous environment within organisms, acids dissociate into hydrogen ions (H^+) and anions. Bases, in contrast, dissociate into hydroxide ions (OH^-) and cations. The more hydrogen ions that are free in a solution, the more acidic the solution is. Conversely, the more hydroxide ions that are free in a solution, the more basic, or alkaline, it is.

Biochemical reactions—that is, chemical reactions in living systems—are extremely sensitive to even small changes in the acidity or alkalinity of the environments in which they occur. In fact, H^+ and OH^- are involved in almost all biochemical processes, and any deviation from a cell's narrow band of normal H^+ and OH^- concentrations can dramatically modify the cell's functions. For this reason, the acids and bases that are continually formed in an organism must be kept in balance.

It is convenient to express the amount of H^+ in a solution by a logarithmic **pH** scale, which ranges from 0 to 14 (Figure 2.7). The term *pH* means potential of hydrogen. On a logarithmic scale, a change of one whole number represents a *tenfold* change from the previous concentration. Thus, a solution of pH 1 has ten times more hydrogen ions than a solution of pH 2 and has 100 times more hydrogen ions than a solution of pH 3.

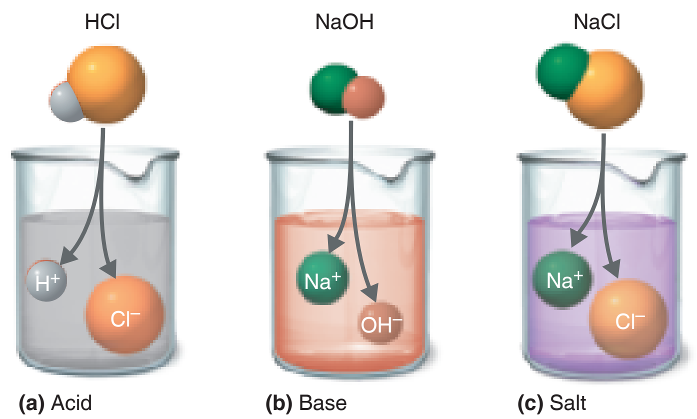


Figure 2.6 Acids, bases, and salts. (a) In water, hydrochloric acid (HCl) dissociates into H^+ and Cl^-. **(b)** Sodium hydroxide (NaOH), a base, dissociates into OH^- and Na^+ in water. **(c)** In water, table salt (NaCl) dissociates into positive ions (Na^+) and negative ions (Cl^-), neither of which are H^+ or OH^-.

Q How do acids and bases differ?

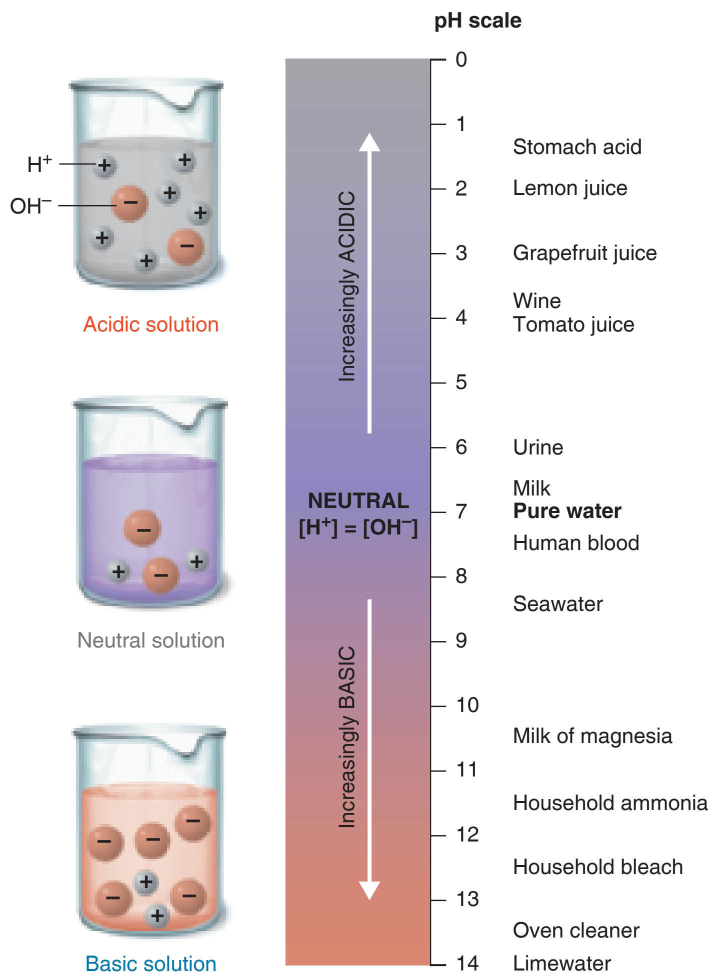


Figure 2.7 The pH scale. As pH values decrease from 14 to 0, the H^+ concentration increases. Thus, the lower the pH, the more acidic the solution; the higher the pH, the more basic the solution. If the pH value of a solution is below 7, the solution is acidic; if the pH is above 7, the solution is basic (alkaline). The approximate pH values of some human body fluids and common substances are shown next to the pH scale.

Q At what pH are the concentrations of H^+ and OH^- equal?

A solution's pH is calculated as $-\log_{10}[H^+]$, the negative logarithm to the base 10 of the hydrogen ion concentration (denoted by brackets), determined in moles per liter $[H^+]$. For example, if the H^+ concentration of a solution is 1.0×10^{-4} moles/liter, or 10^{-4}, its pH equals $-\log_{10}10^{-4} = -(-4) = 4$; this is about the pH value of wine (see Appendix B). The pH values of some human body fluids and other common substances are also shown in Figure 2.7. In the laboratory, you will usually measure the pH of a solution with a pH meter or with chemical test papers.

Acidic solutions contain more H^+ than OH^- and have a pH lower than 7. If a solution has more OH^- than H^+, it is a basic, or alkaline, solution. In pure water, a small percentage of the molecules are dissociated into H^+ and OH^-, so it has a pH of 7. Because the concentrations of H^+ and OH^- are equal, this pH is said to be the pH of a neutral solution.

Keep in mind that the pH of a solution can be changed. We can increase its acidity by adding substances that will increase the concentration of hydrogen ions. As a living organism takes up nutrients, carries out chemical reactions, and excretes wastes, its balance of acids and bases tends to change, and the pH fluctuates. Fortunately, organisms possess natural pH **buffers**, compounds that help keep the pH from changing drastically. But the pH in our environment's water and soil can be altered by waste products from organisms, pollutants from industry, or fertilizers used in agricultural fields or gardens. When bacteria are grown in a laboratory medium, they excrete waste products such as acids that can alter the pH of the medium. If this effect were to continue, the medium would become acidic enough to inhibit bacterial enzymes and kill the bacteria. To prevent this problem, pH buffers are added to the culture medium. One very effective pH buffer for some culture media uses a mixture of K_2HPO_4 and KH_2PO_4 (see Table 6.2, page 159).

Different microbes function best within different pH ranges, but most organisms grow best in environments with a pH value between 6.5 and 8.5. Among microbes, fungi are best able to tolerate acidic conditions, whereas the prokaryotes called cyanobacteria tend to do well in alkaline habitats. *Propionibacterium acnes* (prō-pē′on-ē-bak-TI-rē-um AK-nēz), a bacterium that causes acne, has as its natural environment human skin, which tends to be slightly acidic, with a pH of about 4. *Acidithiobacillus ferrooxidans* (a′sid-ē-thī′ō-bah-SIL-lus fer′rō-OKS-i-danz) is a bacterium that metabolizes elemental sulfur and produces sulfuric acid (H_2SO_4). Its pH range for optimum growth is from 1 to 3.5. The sulfuric acid produced by this bacterium in mine water is important in dissolving uranium and copper from low-grade ore (see Chapter 28).

CHECK YOUR UNDERSTANDING

- **2-4** Why is the polarity of a water molecule important?
- **2-5** Antacids neutralize acid by the following reaction.

$$Mg(OH)_2 + 2HCl \rightarrow MgCl_2 + H_2O$$

Identify the acid, base, and salt.

Organic Compounds

LEARNING OBJECTIVES

- **2-6** Distinguish organic and inorganic compounds.
- **2-7** Define *functional group*.
- **2-8** Identify the building blocks of carbohydrates.
- **2-9** Differentiate simple lipids, complex lipids, and steroids.
- **2-10** Identify the building blocks and structure of proteins.
- **2-11** Identify the building blocks of nucleic acids.
- **2-12** Describe the role of ATP in cellular activities.

Inorganic compounds, excluding water, constitute about 1–1.5% of living cells. These relatively simple components, whose molecules have only a few atoms, cannot be used by cells to perform complex biological functions. Organic molecules, whose carbon atoms can combine in an enormous variety of ways with other carbon atoms and with atoms of other elements, are relatively complex and thus are capable of more complex biological functions.

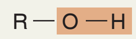
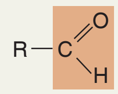
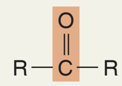
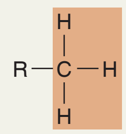
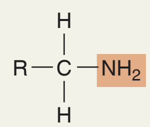
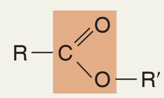
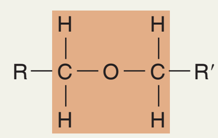
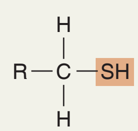
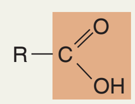
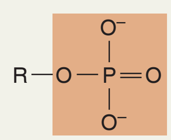
Structure and Chemistry

In the formation of organic molecules, carbon's four outer electrons can participate in up to four covalent bonds, and carbon atoms can bond to each other to form straight-chain, branched-chain, or ring structures.

In addition to carbon, the most common elements in organic compounds are hydrogen (which can form one bond), oxygen (two bonds), and nitrogen (three bonds). Sulfur (two bonds) and phosphorus (five bonds) appear less often. Other elements are found, but only in relatively few organic compounds. The elements that are most abundant in living organisms are the same as those that are most abundant in organic compounds (see Table 2.1).

The chain of carbon atoms in an organic molecule is called the **carbon skeleton**; a huge number of combinations is possible for carbon skeletons. Most of these carbons are bonded to hydrogen atoms. The bonding of other elements with carbon and hydrogen forms characteristic **functional groups**, specific groups of atoms that are most commonly involved in chemical reactions and are responsible for most of the characteristic chemical properties and many of the physical properties of a particular organic compound (Table 2.4).

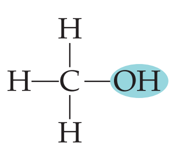
TABLE 2.4 Representative Functional Groups and the Compounds in Which They Are Found

Structure	Name of Group	Biological Importance
	Alcohol	Lipids; carbohydrates
	Aldehyde*	Reducing sugars such as glucose; polysaccharides
	Ketone*	Metabolic intermediates
	Methyl	DNA; energy metabolism
	Amino	Proteins
	Ester	Bacterial and eukaryotic plasma membranes
	Ether	Archaeal plasma membranes
	Sulfhydryl	Energy metabolism; protein structure
	Carboxyl	Organic acids; lipids; proteins
	Phosphate	ATP; DNA

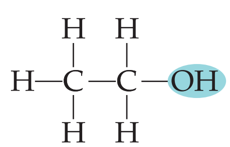
*In an aldehyde, a C═O is at the end of a molecule, in contrast to the internal C═O in a ketone.

Different functional groups confer different properties on organic molecules. For example, the hydroxyl group of alcohols is hydrophilic (water-loving) and thus attracts water molecules to it. This attraction helps dissolve organic molecules containing hydroxyl groups. Because the carboxyl group is a source of hydrogen ions, molecules containing it have acidic properties. Amino groups, by contrast, function as bases because they readily accept hydrogen ions. The sulfhydryl group helps stabilize the intricate structure of many proteins.

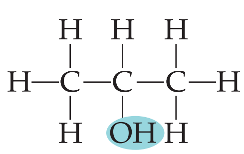
Functional groups help us classify organic compounds. For example, the —OH group is present in each of the following molecules:



Methanol



Ethanol

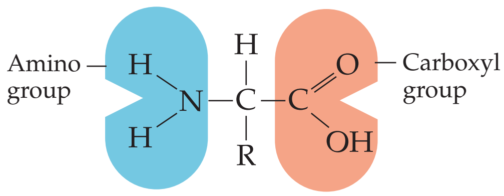


Isopropanol

Because the characteristic reactivity of the molecules is based on the —OH group, they are grouped together in a class called alcohols. The —OH group is called the *hydroxyl group* and is not to be confused with the *hydroxide ion* (OH^-) of bases. The hydroxyl group of alcohols does not ionize at neutral pH; it is covalently bonded to a carbon atom.

When a class of compounds is characterized by a certain functional group, the letter *R* can be used to stand for the remainder of the molecule. For example, alcohols in general may be written R—OH.

Frequently, more than one functional group is found in a single molecule. For example, an amino acid molecule contains both amino and carboxyl groups. The amino acid glycine has the following structure:



Most of the organic compounds found in living organisms are quite complex; a large number of carbon atoms form the skeleton, and many functional groups are attached. In organic molecules, it is important that each of the four bonds of carbon be satisfied (attached to another atom) and that each of the attaching atoms have its characteristic number of bonds satisfied. Because of this, such molecules are chemically stable.

Small organic molecules can be combined into very large molecules called **macromolecules** (*macro* = large). Macromolecules are usually **polymers** (*poly* = many; *mers* = parts): polymers are formed by covalent bonding of many repeating small molecules called **monomers** (*mono* = one). When two monomers join together, the reaction usually involves the elimination of a hydrogen atom from one monomer and a hydroxyl group from the other; the hydrogen atom and the hydroxyl group combine to produce water:

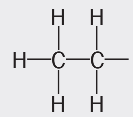
$$R\text{—}OH + OH\text{—}R' \longrightarrow R\text{—}R' + H_2O$$

This type of exchange reaction is called **dehydration synthesis** (*de* = from; *hydro* = water), or a **condensation reaction**, because a molecule of water is released (**Figure 2.8a**). Such macromolecules as carbohydrates, lipids, proteins, and nucleic acids are assembled in the cell, essentially by dehydration synthesis. However, other molecules must also participate to provide energy for bond formation. ATP, the cell's chief energy provider, is discussed at the end of this chapter.

CHECK YOUR UNDERSTANDING

✓ **2-6** Define *organic*.

✓ **2-7** Add the appropriate functional group(s) to the ethyl group below to produce each of the following compounds: ethanol, acetic acid, acetaldehyde, ethanolamine, diethyl ether.



Carbohydrates

The **carbohydrates** are a large and diverse group of organic compounds that includes sugars and starches. Carbohydrates perform a number of major functions in living systems. For instance, one type of sugar (deoxyribose) is a building block of deoxyribonucleic acid (DNA), the molecule that carries hereditary information. Other sugars are needed for the cell walls. Simple carbohydrates are used in the synthesis of amino acids and fats or fatlike substances, which are used to build cell membranes and other structures. Macromolecular carbohydrates function as food reserves. The principal function of carbohydrates, however, is to fuel cell activities with a ready source of energy.

Carbohydrates are made up of carbon, hydrogen, and oxygen atoms. The ratio of hydrogen to oxygen atoms is always 2:1 in simple carbohydrates. This ratio can be seen in the formulas for the carbohydrates ribose ($C_5H_{10}O_5$), glucose ($C_6H_{12}O_6$), and sucrose ($C_{12}H_{22}O_{11}$). Although there are exceptions, the

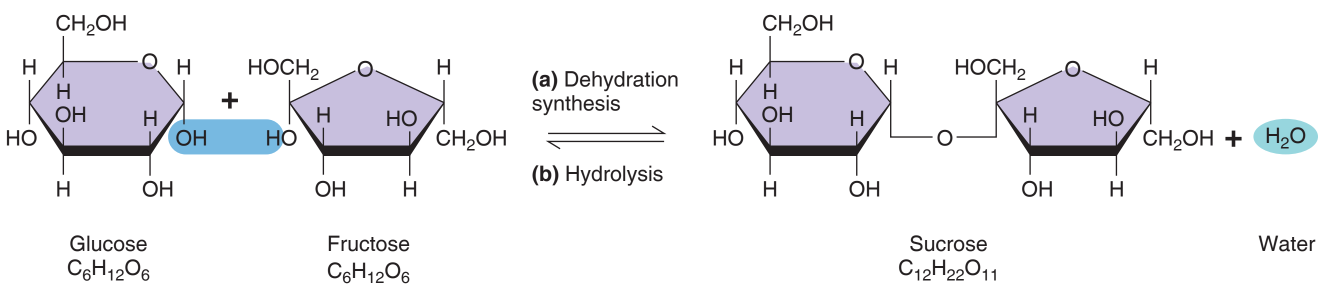


Figure 2.8 Dehydration synthesis and hydrolysis. (a) In dehydration synthesis (left to right), the monosaccharides glucose and fructose combine to form a molecule of the disaccharide sucrose. A molecule of water is released in the reaction. **(b)** In hydrolysis (right to left), the sucrose molecule breaks down into the smaller molecules glucose and fructose. For the hydrolysis reaction to proceed, water must be added to the sucrose.

Q What is the difference between a polymer and a monomer?

general formula for carbohydrates is $(CH_2O)_n$, where n indicates that there are three or more CH_2O units. Carbohydrates can be classified into three major groups on the basis of size: monosaccharides, disaccharides, and polysaccharides.

Monosaccharides

Simple sugars are called **monosaccharides** (*sacchar* = sugar); each molecule contains three to seven carbon atoms. The number of carbon atoms in the molecule of a simple sugar is indicated by the prefix in its name. For example, simple sugars with three carbons are called trioses. There are also tetroses (four-carbon sugars), pentoses (five-carbon sugars), hexoses (six-carbon sugars), and heptoses (seven-carbon sugars). Pentoses and hexoses are extremely important to living organisms. Deoxyribose is a pentose found in DNA. Glucose, a very common hexose, is the main energy-supplying molecule of living cells.

Disaccharides

Disaccharides (*di* = two) are formed when two monosaccharides bond in a dehydration synthesis reaction.* For example, molecules of two monosaccharides, glucose and fructose, combine to form a molecule of the disaccharide sucrose (table sugar) and a molecule of water (see Figure 2.8a). Similarly, the dehydration synthesis of the monosaccharides glucose and galactose forms the disaccharide lactose (milk sugar).

It may seem odd that glucose and fructose have the same chemical formula (see Figure 2.8), even though they are different monosaccharides. The positions of the oxygens and carbons differ in the two different molecules; consequently, the molecules have different physical and chemical properties. Two molecules with the same chemical formula but different structures and properties are called **isomers** (*iso* = same).

Disaccharides can be broken down into smaller, simpler molecules when water is added. This chemical reaction, the reverse of dehydration synthesis, is called **hydrolysis** (*hydro* = water; *lysis* = to loosen) (Figure 2.8b). A molecule of sucrose, for example, may be hydrolyzed (digested) into its components of glucose and fructose by reacting with the H^+ and OH^- of water.

As you will see in Chapter 4, the cell walls of bacterial cells are composed of disaccharides and proteins, which together are called peptidoglycan.

Polysaccharides

Carbohydrates in the third major group, the **polysaccharides**, consist of tens or hundreds of monosaccharides joined through dehydration synthesis. Polysaccharides often have side chains branching off the main structure and are classified as macromolecules. Like disaccharides, polysaccharides can be split apart into their constituent sugars through hydrolysis. Unlike monosaccharides and disaccharides, however, they usually lack the characteristic sweetness of sugars such as fructose and sucrose and usually are not soluble in water.

One important polysaccharide is *glycogen*, which is composed of glucose subunits and is synthesized as a storage material by animals and some bacteria. *Cellulose*, another important glucose polymer, is the main component of the cell walls of plants and most algae. Although cellulose is the most abundant carbohydrate on Earth, it can be digested by only a few organisms that have the appropriate enzyme. The polysaccharide *dextran*, which is produced as a sugary slime by certain bacteria, is used in a blood plasma substitute. *Chitin* is a polysaccharide that makes up part of the cell wall of most fungi and the exoskeletons of lobsters, crabs, and insects. *Starch* is a polymer of glucose produced by plants and used as food by humans. Digestion of starch by intestinal bacteria is important to human health (see Exploring the Microbiome, page 37).

Many animals, including humans, produce enzymes called *amylases* that can break the bonds between the glucose molecules in glycogen. However, this enzyme cannot break the bonds in cellulose. Bacteria and fungi that produce enzymes called *cellulases* can digest cellulose. Cellulases from the fungus *Trichoderma* (TRIK-ō-der-mah) are used for a variety of industrial purposes. One of the more unusual uses is producing stone-washed or distressed denim. Because washing the fabric with rocks would damage washing machines, cellulase is used to digest, and therefore soften, the cotton.

CHECK YOUR UNDERSTANDING

✓ **2-8** Give an example of a monosaccharide, a disaccharide, and a polysaccharide.

Lipids

If lipids were suddenly to disappear from the Earth, all living cells would collapse in a pool of fluid, because lipids are essential to the structure and function of membranes that separate living cells from their environment. **Lipids** (*lip* = fat) are a second major group of organic compounds found in living matter. Like carbohydrates, they are composed of atoms of carbon, hydrogen, and oxygen, but lipids lack the 2:1 ratio between hydrogen and oxygen atoms. Even though lipids are a very diverse group of compounds, they share one common characteristic: they are *nonpolar* molecules so, unlike water, do not have a positive and a negative end (pole). Therefore, most lipids are insoluble in water but dissolve readily in nonpolar solvents, such as ether and chloroform. Lipids provide the structure of membranes and some cell walls and function in energy storage.

*Carbohydrates composed of 2 to about 20 monosaccharides are called **oligosaccharides** (*oligo* = few). Disaccharides are the most common oligosaccharides.

EXPLORING THE MICROBIOME Feed Our Intestinal Bacteria, Feed Ourselves: A Tale of Two Starches

Structurally speaking, starch found in plants appears as either a many-branched chain called amylopectin or as a straight-chained variety called amylose. Amylopectin-rich foods common to our diet include sticky rice and waxy varieties of corn and potatoes. Foods containing high levels of amylose include beans and other legumes and whole grains. Although both are starches, they produce markedly different effects when we eat them.

In the small intestine, enzymes rapidly convert amylopectin into glucose, which is the preferred carbohydrate for many essential metabolic reactions our cells conduct. By contrast, amylose's structure has less surface area for enzymes to react with, making it more resistant to digestion. Since it isn't easily broken down in our small intestine, the amylose starch continues through the gastrointestinal tract to the colon, where it's available for bacteria living there to ferment.

You might assume that amylopectin, the starch that is easily broken down, must be the best type for us to eat. However, feeding our microbiome amylose seems to provide excellent health benefits. Amylose-fermenting bacteria, including members of the *Prevotella* and *Lachnospira* genera, produce short-chain fatty acids. Several types of these molecules are important players in how our intestinal cells absorb electrolytes (ions).

Research also suggests that butyrate, one type of short-chain fatty acid linked to *Prevotella* metabolism, may protect us against colorectal cancer. In another study, mice were treated with antibiotics and then inoculated with pathogenic *Clostridium difficile* bacteria. What happened to the animals next was stark: some rapidly developed lethal infections, while others were colonized by the bacteria and experienced only mild disease. The fate of the study mice seemed to come down to their microbiome composition before *C. difficile* was introduced. Mice with large numbers of *Lachnospira* bacteria were more likely to survive, whereas those with intestinal microbiomes dominated by *E. coli* were much more likely to die.

Helpful *Prevotella* species thrive in the intestines by fermenting amylose we can't digest.

Simple Lipids

Simple lipids, called *fats* or *triglycerides*, contain an alcohol called *glycerol* and a group of compounds known as *fatty acids*. Glycerol molecules have three carbon atoms to which are attached three hydroxyl (—OH) groups (**Figure 2.9a**). Fatty acids consist of long hydrocarbon chains (composed only of carbon and hydrogen atoms) ending in a carboxyl (—COOH, organic acid) group (Figure 2.9b). Most common fatty acids contain an even number of carbon atoms.

A molecule of fat is formed when a molecule of glycerol combines with one to three fatty acid molecules. The number of fatty acid molecules determines whether the fat molecule is a monoglyceride, diglyceride, or triglyceride (Figure 2.9c). In the reaction, one to three molecules of water are formed (dehydration), depending on the number of fatty acid molecules reacting. The chemical bond formed where the water molecule is removed is called an *ester linkage*. In the reverse reaction, hydrolysis, a fat molecule is broken down into its component fatty acid and glycerol molecules.

Because the fatty acids that form lipids have different structures, there is a wide variety of lipids. For example, three molecules of fatty acid A might combine with a glycerol molecule. Or one molecule each of fatty acids A, B, and C might unite with a glycerol molecule (see Figure 2.9c).

The primary function of lipids is to form plasma membranes that enclose cells. A plasma membrane supports the cell and allows nutrients and wastes to pass in and out; therefore, the lipids must maintain the same viscosity, regardless of the surrounding temperature. The membrane must be about as viscous as olive oil, without getting too fluid when warmed or too thick when cooled. As everyone who has ever cooked a meal knows, animal fats (such as butter) are usually solid at room temperature, whereas vegetable oils are usually liquid at room temperature. The difference in their respective melting points is due to the degrees of saturation of the fatty acid chains. A fatty acid is said to be *saturated* when it has no double bonds, in which case the carbon skeleton contains the maximum

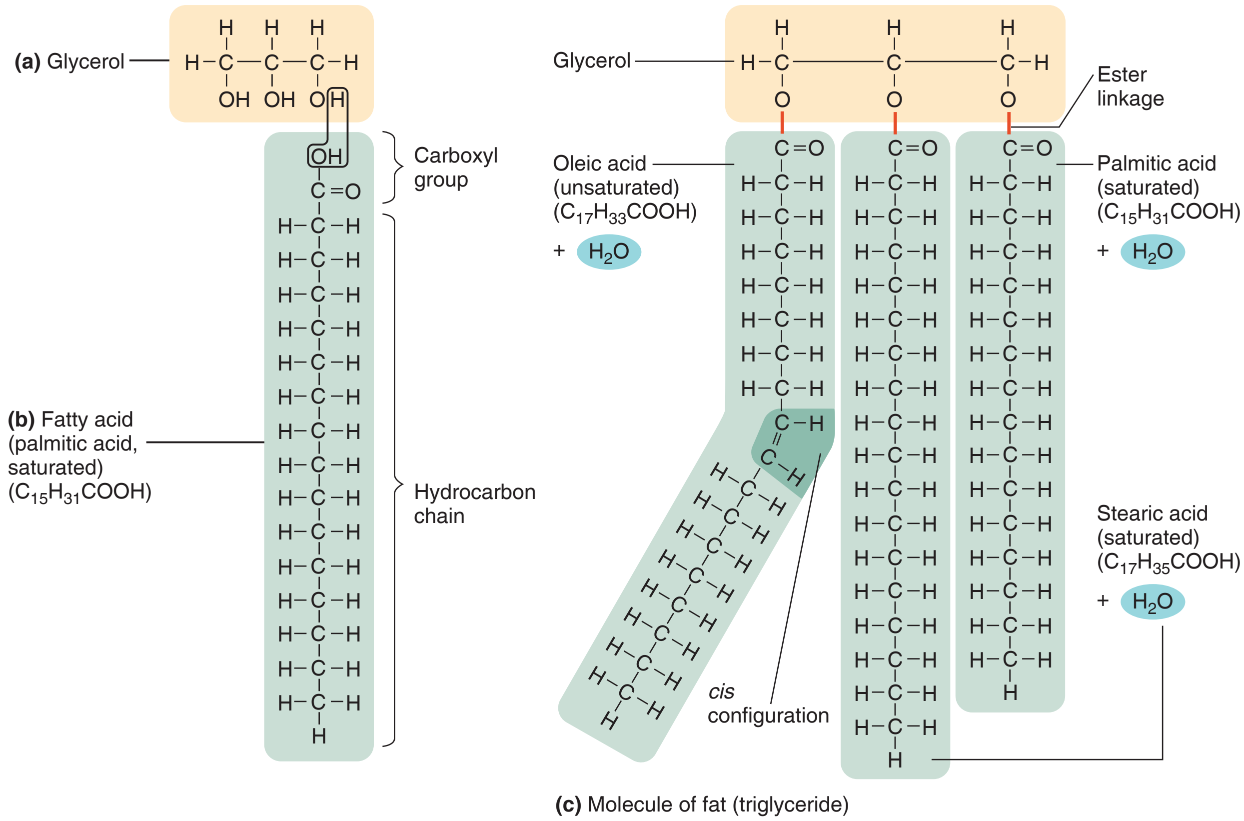


Figure 2.9 Structural formulas of simple lipids. (a) Glycerol. **(b)** Palmitic acid, a saturated fatty acid. **(c)** The chemical combination of a molecule of glycerol and three fatty acid molecules (palmitic, stearic, and oleic in this example) forms one molecule of fat (triglyceride) and three molecules of water in a dehydration synthesis reaction. Oleic acid is a *cis* fatty acid. The bond between glycerol and each fatty acid is called an ester linkage. The addition of three water molecules to a fat forms glycerol and three fatty acid molecules in a hydrolysis reaction.

Q How do saturated and unsaturated fatty acids differ?

number of hydrogen atoms (see Figure 2.9c and **Figure 2.10a**). Saturated chains become solid more easily because they are relatively straight and are thus able to pack together more closely than unsaturated chains. The double bonds of *unsaturated* chains create kinks in the chain, which keep the chains apart from one another (Figure 2.10b). Note in Figure 2.9c that the H atoms on either side of the double bond in oleic acid are on the same side of the unsaturated fatty acid. Such an unsaturated fatty acid is called a *cis* fatty acid. If, instead, the H atoms are on opposite sides of the double bond, the unsaturated acid is called a *trans* fatty acid.

Complex Lipids

Complex lipids contain such elements as phosphorus, nitrogen, and sulfur, in addition to the carbon, hydrogen, and oxygen found in simple lipids. The complex lipids called *phospholipids* are made up of glycerol, two fatty acids, and, in place of a third fatty acid, a phosphate group bonded to one of several organic groups (see Figure 2.10a). Phospholipids are the lipids that build membranes; they are essential to a cell's survival. Phospholipids have polar as well as nonpolar regions (Figure 2.10a and b; see also Figure 4.14, page 86). When placed in water, phospholipid molecules twist themselves in such a way that all polar (hydrophilic) portions orient themselves toward the polar water molecules, with which they then form hydrogen bonds. (Recall that *hydrophilic* means water-loving.) This forms the basic structure of a plasma membrane (Figure 2.10c). Polar portions consist of a phosphate group and glycerol. In contrast to the polar regions, all nonpolar (hydrophobic) parts of the phospholipid make contact only with the nonpolar portions of neighboring molecules. (*Hydrophobic* means water-fearing.) Nonpolar portions consist of fatty acids. This characteristic behavior makes phospholipids particularly suitable for their role as a major component of the membranes that enclose cells. Phospholipids enable the membrane to act as a barrier that separates the contents of the cell from the water-based environment in which it lives.

Some complex lipids are useful in identifying certain bacteria. For example, the cell wall of *Mycobacterium tuberculosis* (mī′kō-bak-TI-rē-um too′ber-kū-LŌ-sis), the bacterium that causes tuberculosis, is distinguished by its lipid-rich content. The cell wall contains complex lipids such as waxes and glycolipids (lipids

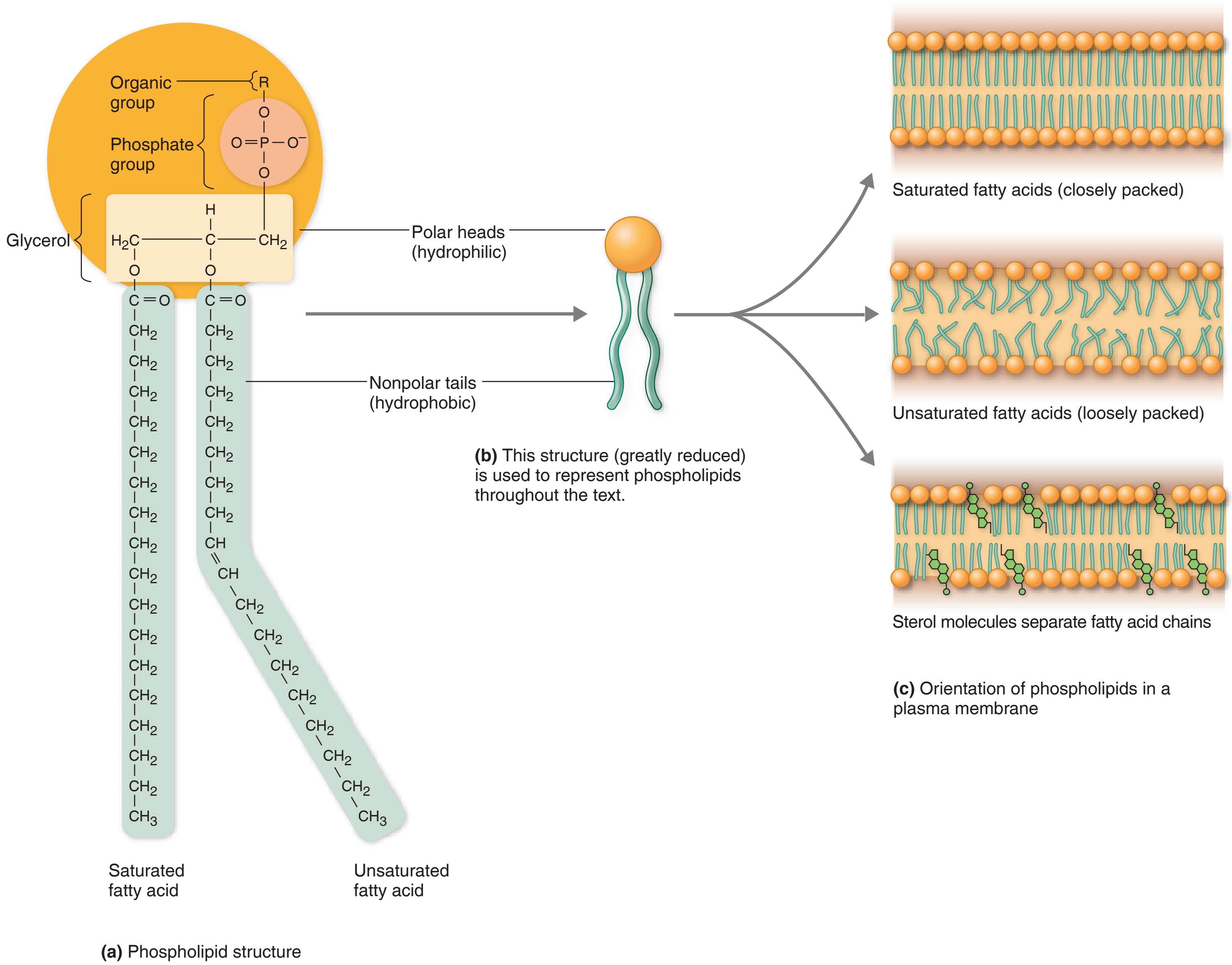


Figure 2.10 Phospholipid and orientation, showing saturated and unsaturated fatty acids and the molecules' polarity.

Q Where are phospholipids found in cells?

with carbohydrates attached) that give the bacterium distinctive staining characteristics. Cell walls rich in such complex lipids are characteristic of all members of the genus *Mycobacterium*.

Steroids

Steroids are structurally very different from lipids. Figure 2.11 shows the structure of the steroid cholesterol, with the four interconnected carbon rings that are characteristic of steroids. When an —OH group is attached to one of the rings, the steroid is called a *sterol* (an alcohol). Sterols are important constituents of the plasma membranes of animal cells and of one group of bacteria (mycoplasmas), and they are also found in fungi and plants. The sterols separate the fatty acid chains and

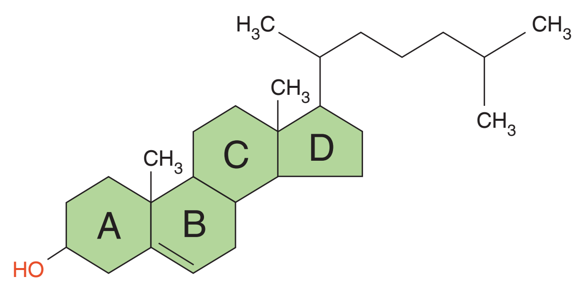


Figure 2.11 Cholesterol, a steroid. Note the four "fused" carbon rings (labeled A–D), which are characteristic of steroid molecules. The hydrogen atoms attached to the carbons at the corners of the rings have been omitted. The —OH group (colored red) makes this molecule a sterol.

Q Where are sterols found in cells?

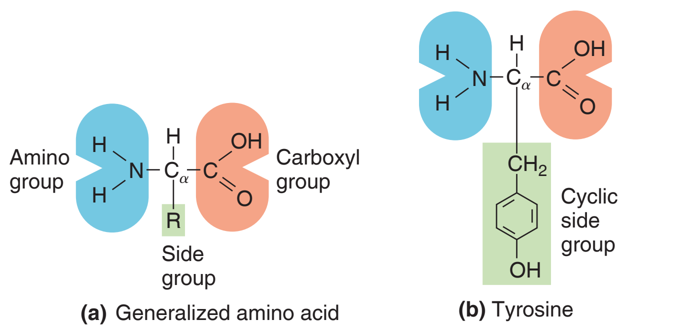


Figure 2.12 Amino acid structure. **(a)** The general structural formula for an amino acid. The alpha-carbon (C_α) is shown in the center. Different amino acids have different R groups, also called side groups. **(b)** Structural formula for the amino acid tyrosine, which has a cyclic side group.

Q What distinguishes one amino acid from another?

thus prevent the packing that would harden the plasma membrane at low temperatures (see Figure 2.10c).

CHECK YOUR UNDERSTANDING

✓ **2-9** How do simple lipids differ from complex lipids?

Proteins

Proteins are organic molecules that contain carbon, hydrogen, oxygen, and nitrogen. Some also contain sulfur. If you were to separate and weigh all the groups of organic compounds in a living cell, the proteins would tip the scale. Hundreds of different proteins can be found in any single cell, and together they make up 50% or more of a cell's dry weight.

Proteins are essential ingredients in all aspects of cell structure and function. *Enzymes* are the proteins that speed up biochemical reactions. But proteins have other functions as well. *Transporter proteins* help transport certain chemicals into and out of cells. Other proteins, such as the *bacteriocins* produced by many bacteria, kill other bacteria. Certain *toxins*, called exotoxins, produced by some disease-causing microorganisms are also proteins. Some proteins play a role in the *contraction* of animal muscle cells and the *movement* of microbial and other types of cells. Other proteins are integral parts of *cell structures* such as walls, membranes, and cytoplasmic components. Still others, such as the *hormones* of certain organisms, have regulatory functions. As we will see in Chapter 17, proteins called *antibodies* play a role in vertebrate immune systems.

Amino Acids

Just as monosaccharides are the building blocks of larger carbohydrate molecules, and just as fatty acids and glycerol are the building blocks of fats, **amino acids** are the building blocks of proteins. Amino acids contain at least one carboxyl (—COOH) group and one amino (—NH_2) group attached to the same carbon atom, called an alpha-carbon (written C_α) (Figure 2.12a). Such amino acids are called *alpha-amino acids*. Also attached to the alpha-carbon is a side group (R group), which is the amino acid's distinguishing feature. The side group can be a hydrogen atom, an unbranched or branched chain of atoms, or a ring structure that is cyclic (all carbon) or heterocyclic (when an atom other than carbon is included in the ring). Figure 2.12b shows the structural formula of tyrosine, an amino acid that has a cyclic side group. The side group can contain functional groups, such as the sulfhydryl group (—SH), the hydroxyl group (—OH), or additional carboxyl or amino groups. These side groups and the carboxyl and alpha-amino groups affect the total structure of a protein, described later. The structures and standard abbreviations of the 20 amino acids found in proteins are shown in Table 2.5.

Most amino acids exist in either of two configurations called **stereoisomers**, designated by D and L. These configurations are mirror images, corresponding to "right-handed" (D) and "left-handed" (L) three-dimensional shapes (Figure 2.13). The amino acids found in proteins are always the L-isomers (except for glycine, the simplest amino acid, which does not

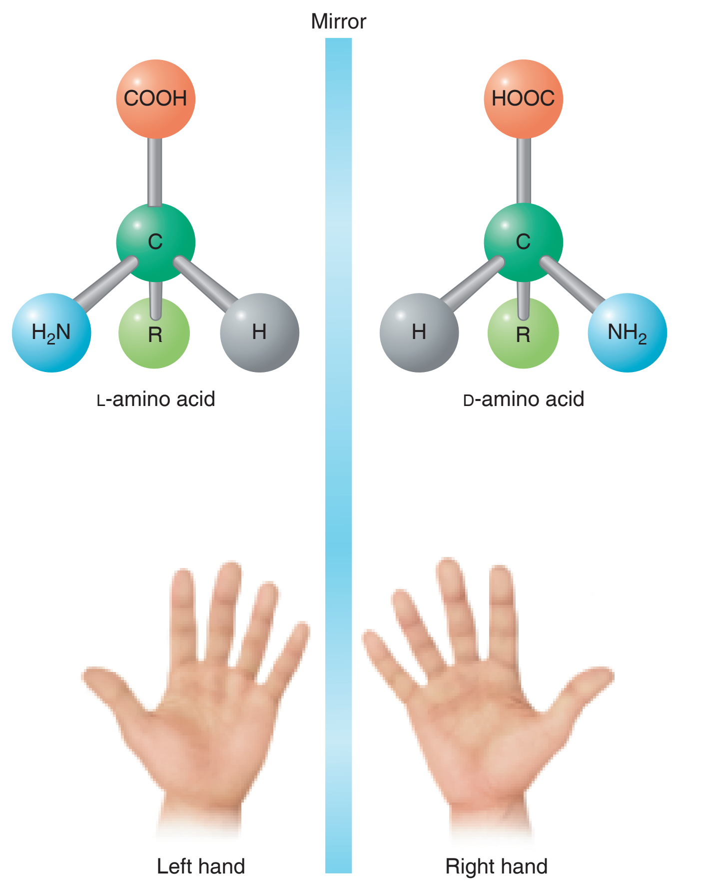
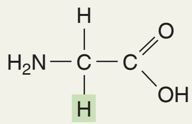
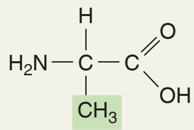
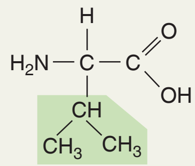
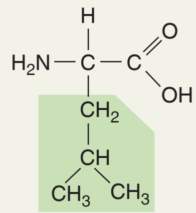
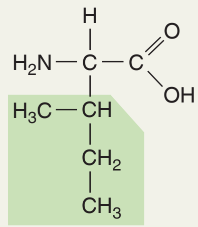
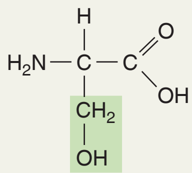
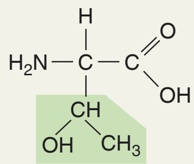
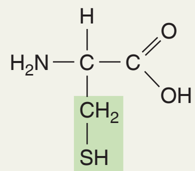
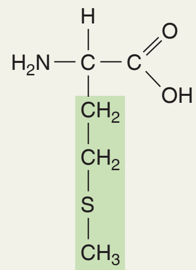
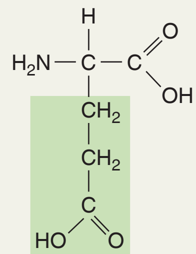
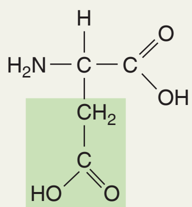
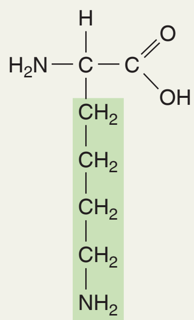
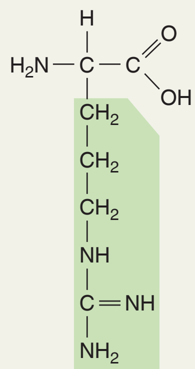
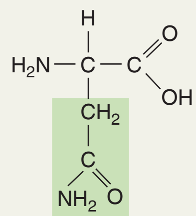
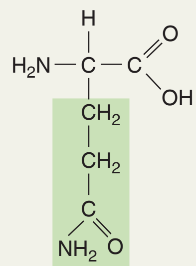
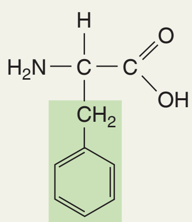
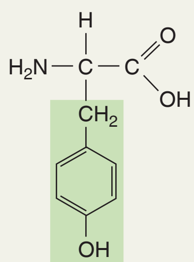
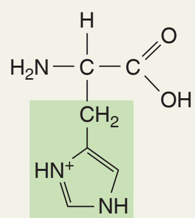
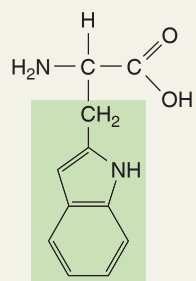
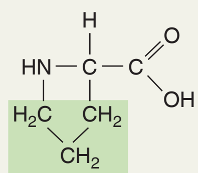


Figure 2.13 The L- and D-isomers of an amino acid, shown with ball-and-stick models. The two isomers, like left and right hands, are mirror images of each other and cannot be superimposed on one another. (Try it!)

Q Which isomer is always found in proteins?

TABLE 2.5 The 20 Amino Acids Found in Proteins*

Glycine (Gly, G)	Alanine (Ala, A)	Valine (Val, V)	Leucine (Leu, L)	Isoleucine (Ile, I)
				
Hydrogen atom	Unbranched chain	Branched chain	Branched chain	Branched chain
Serine (Ser, S)	**Threonine (Thr, T)**	**Cysteine (Cys, C)**	**Methionine (Met, M)**	**Glutamic acid (Glu, E)**
				
Hydroxyl (—OH) group	Hydroxyl (—OH) group	Sulphur-containing (—SH) group	Thioether (SC) group	Additional carboxyl (—COOH) group, acidic
Aspartic acid (Asp, D)	**Lysine (Lys, K)**	**Arginine (Arg, R)**	**Asparagine (Asn, N)**	**Glutamine (Gln, Q)**
				
Additional Carboxyl (—COOH) group, acidic	Additional amino (—NH_2) group, basic	Additional amino (—NH_2) group, basic	Additional amino (—NH_2) group, basic	Additional amino (—NH_2) group, basic
Phenylalanine (Phe, F)	**Tyrosine (Tyr, Y)**	**Histidine (His, H)**	**Tryptophan (Trp, W)**	**Proline (Pro, P)**
				
Cyclic	Cyclic	Heterocyclic	Heterocyclic	Heterocyclic

*Shown are the amino acid names, including the three-letter and one-letter abbreviations in parentheses (above), their structural formulas (center), and characteristic R group (in green). Note that cysteine and methionine are the only amino acids that contain sulfur.

CLINICAL CASE

While Jonathan is in intensive care, his wife, DeeAnn, and adult daughter talk with his physician and an investigator from the Centers for Disease Control and Prevention (CDC) to find the source of Jonathan's *B. anthracis* infection. Environmental investigations uncover *B. anthracis* at Jonathan's home, in his van, and in his workplace, but neither his wife nor children show signs of infection. His bandmates are also tested; they are all negative for *B. anthracis*. The CDC investigator explains to Jonathan's family that *B. anthracis* forms endospores that can survive in soil for up to 60 years. It is rare in humans; however, grazing animals and people who handle their hides or other by-products can become infected. *B. anthracis* cells have capsules that are composed of poly-D-glutamic acid.

Why are the capsules resistant to digestion by phagocytes? (Phagocytes are white blood cells that engulf and destroy bacteria.)

25 **42** 44 46

have stereoisomers). However, D-amino acids occasionally occur in nature—for example, in certain bacterial cell walls and antibiotics.

Although only 20 different amino acids occur naturally in proteins, a single protein molecule can contain from 50 to hundreds of amino acid molecules, which can be arranged in an almost infinite number of ways to make proteins of different lengths, compositions, and structures. The number of proteins is practically endless, and every living cell produces many different proteins.

Peptide Bonds

Amino acids bond between the carbon atom of the carboxyl (—COOH) group of one amino acid and the nitrogen atom of the amino (—NH_2) group of another (**Figure 2.14**). The bonds between amino acids are called **peptide bonds.** For every peptide bond formed between two amino acids, one water molecule is released; thus, peptide bonds are formed by dehydration synthesis. The resulting compound in Figure 2.14 is called a *dipeptide* because it consists of two amino acids joined by a peptide bond. Adding another amino acid to a dipeptide would form a *tripeptide*. Further additions of amino acids would produce a long, chainlike molecule called a *peptide* (4–9 amino acids) or *polypeptide* (10–2000 or more amino acids).

Levels of Protein Structure

Proteins vary tremendously in structure. Different proteins have different architectures and different three-dimensional shapes. This variation in structure is directly related to their diverse functions.

When a cell makes a protein, the polypeptide chain folds spontaneously to assume a certain shape. One reason for folding of the polypeptide is that some parts of a protein are attracted to water and other parts are repelled by it. In practically every case, the function of a protein depends on its ability to recognize and bind to some other molecule. For example, an enzyme binds specifically with its substrate. A hormonal protein binds to a receptor on a cell whose function it will alter. An antibody binds to an antigen (foreign substance) that has invaded the body. The unique shape of each protein permits it to interact with specific other molecules in order to carry out specific functions.

Proteins are described in terms of four levels of organization: primary, secondary, tertiary, and quaternary. The *primary structure* is the unique sequence in which the amino acids are linked together to form a polypeptide chain (**Figure 2.15a**). This sequence is genetically determined. Alterations in sequence can have profound metabolic effects. For example, a single incorrect amino acid in a blood protein can produce the deformed hemoglobin molecule characteristic of sickle cell disease. But proteins do not exist as long, straight chains. Each polypeptide chain folds and coils in specific ways into a relatively compact structure with a characteristic three-dimensional shape.

A protein's *secondary structure* is the localized, repetitious twisting or folding of the polypeptide chain. This aspect of a protein's shape results from hydrogen bonds joining the atoms of peptide bonds at different locations along the polypeptide chain. The two types of secondary protein structures are clockwise spirals called *helices* (singular: *helix*) and pleated sheets, which form from roughly parallel portions of the chain (Figure 2.15b). Both structures are held together by hydrogen bonds between oxygen or nitrogen atoms that are part of the polypeptide's backbone.

Tertiary structure refers to the overall three-dimensional structure of a polypeptide chain (Figure 2.15c). The folding is not repetitive or predictable, as in secondary structure. Whereas secondary structure involves hydrogen bonding between atoms

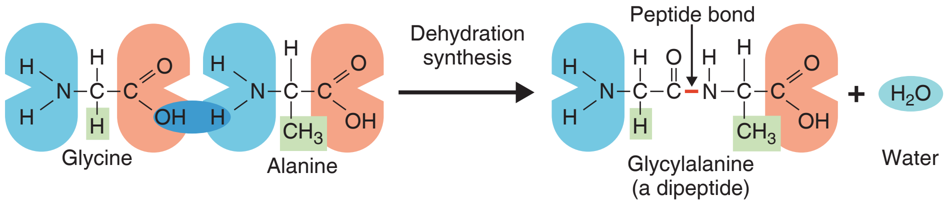


Figure 2.14 Peptide bond formation by dehydration synthesis. The amino acids glycine and alanine combine to form a dipeptide. The newly formed bond between the carbon atom of glycine and the nitrogen atom of alanine is called a peptide bond.

Q How are amino acids related to proteins?

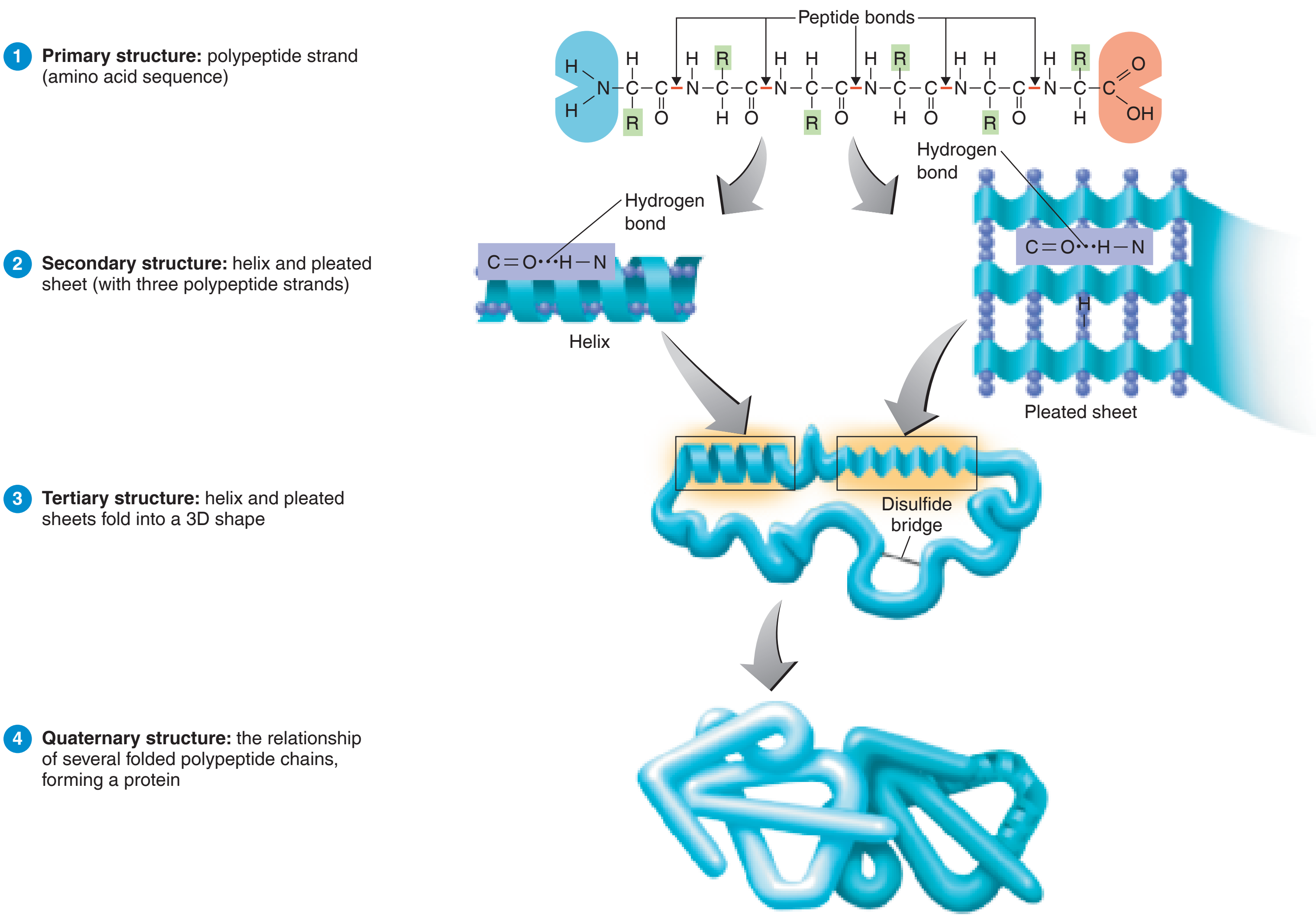


Figure 2.15 Protein structure. 1 Primary structure, the amino acid sequence. 2 Secondary structures: helix and pleated sheet. 3 Tertiary structure, the overall three-dimensional folding of a polypeptide chain. 4 Quaternary structure, the relationship between several polypeptide chains that make up a protein. Shown here is the quaternary structure of a hypothetical protein composed of two polypeptide chains.

Q What property of a protein enables it to carry out specific functions?

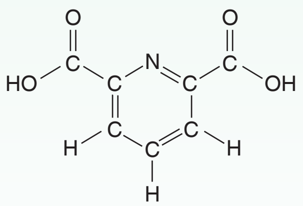
of the amino and carboxyl groups involved in the peptide bonds, tertiary structure involves several interactions between various amino acid side groups in the polypeptide chain. For example, amino acids with nonpolar (hydrophobic) side groups usually interact at the core of the protein, out of contact with water. This *hydrophobic interaction* helps contribute to tertiary structure. Hydrogen bonds between side groups, and ionic bonds between oppositely charged side groups, also contribute to tertiary structure. Proteins that contain the amino acid cysteine form strong covalent bonds called *disulfide bridges*. These bridges form when two cysteine molecules are brought close together by the folding of the protein. Cysteine molecules contain sulfhydryl groups (—SH), and the sulfur of one cysteine molecule bonds to the sulfur on another, forming (by the removal of hydrogen atoms) a disulfide bridge (S—S) that holds parts of the protein together.

Some proteins have a *quaternary structure*, which consists of an aggregation of two or more individual polypeptide chains (subunits) that operate as a single functional unit. Figure 2.15d shows a hypothetical protein consisting of two polypeptide chains. More commonly, proteins have two or more kinds of polypeptide subunits. The bonds that hold a quaternary structure together are basically the same as those that maintain tertiary structure. The overall shape of a protein may be globular (compact and roughly spherical) or fibrous (threadlike).

If a protein encounters a hostile environment in terms of temperature, pH, or salt concentrations, it may unravel and lose its characteristic shape. This process is called **denaturation**

CLINICAL CASE

The host's phagocytes cannot easily digest D-forms of amino acids, such as D-glutamic acid found in the capsules of *B. anthracis*. Therefore, infection can develop. The CDC investigator's mention of animal hides gives DeeAnn an idea. Jonathan plays West African drums called *djembe*; the drum skins are made from dried imported goat hides from West Africa. Although most of these hides are legally imported, some slip through the cracks. It's possible that the hides on Jonathan's drums have been illegally imported and therefore have not been inspected by the U.S. Department of Agriculture. To create *djembe* drums, the hides are soaked in water, stretched over the drum body, and then scraped and sanded. The scraping and sanding generates a large amount of aerosolized dust as the hides dry. Sometimes this dust contains *B. anthracis* endospores, which contain dipicolinic acid.



What is the functional group in dipicolinic acid? See the figure above.

25 42 **44** 46

(see Figure 5.6, page 115). As a result of denaturation, the protein is no longer functional. This process will be discussed in more detail in Chapter 5 with regard to denaturation of enzymes.

The proteins we have been discussing are *simple proteins*, which contain only amino acids. *Conjugated proteins* are combinations of amino acids with other organic or inorganic components. Conjugated proteins are named by their non–amino acid component. Thus, glycoproteins contain sugars, nucleoproteins contain nucleic acids, metalloproteins contain metal atoms, lipoproteins contain lipids, and phosphoproteins contain phosphate groups. Phosphoproteins are important regulators of activity in eukaryotic cells. Bacterial synthesis of phosphoproteins may be important for the survival of bacteria such as *Legionella pneumophila* that grow inside host cells.

CHECK YOUR UNDERSTANDING

✓ **2-10** What two functional groups are in all amino acids?

Nucleic Acids

In 1944, three American microbiologists—Oswald Avery, Colin MacLeod, and Maclyn McCarty—discovered that a substance called **deoxyribonucleic acid (DNA)** is the substance of which genes are made. Nine years later, James Watson and Francis Crick, working with molecular models and X-ray information supplied by Maurice Wilkins and Rosalind Franklin, identified the physical structure of DNA. In addition, Crick suggested a mechanism for DNA replication and how it works as the hereditary material. DNA and another substance called **ribonucleic acid (RNA)** are together referred to as **nucleic acids** because they were first discovered in the nuclei of cells. Just as amino acids are the structural units of proteins, nucleotides are the structural units of nucleic acids.

Each **nucleotide** has three parts: a nitrogen-containing base, a pentose (five-carbon) sugar (either **deoxyribose** or **ribose**), and a phosphate group (phosphoric acid). The nitrogen-containing bases are cyclic compounds made up of carbon, hydrogen, oxygen, and nitrogen atoms. The bases are named adenine (A), thymine (T), cytosine (C), guanine (G), and uracil (U). A and G are double-ring structures called **purines**, whereas T, C, and U are single-ring structures referred to as **pyrimidines**.

Nucleotides are named according to their nitrogen-containing base. Thus, a nucleotide containing thymine is a *thymine nucleotide*, one containing adenine is an *adenine nucleotide*, and so on. The term **nucleoside** refers to the combination of a purine or pyrimidine plus a pentose sugar; it does not contain a phosphate group.

DNA

According to the model proposed by Watson and Crick, a DNA molecule consists of two long strands wrapped around each other to form a **double helix** (Figure 2.16). The double helix looks like a twisted ladder, and each strand is composed of many nucleotides.

Every strand of DNA composing the double helix has a "backbone" consisting of alternating deoxyribose sugar and phosphate groups. The deoxyribose of one nucleotide is joined to the phosphate group of the next. The nitrogen-containing bases make up the rungs of the ladder. Note that the purine A is always paired with the pyrimidine T and that the purine G is always paired with the pyrimidine C. The bases are held together by hydrogen bonds; A and T are held by two hydrogen bonds, and G and C by three. DNA does not contain uracil (U).

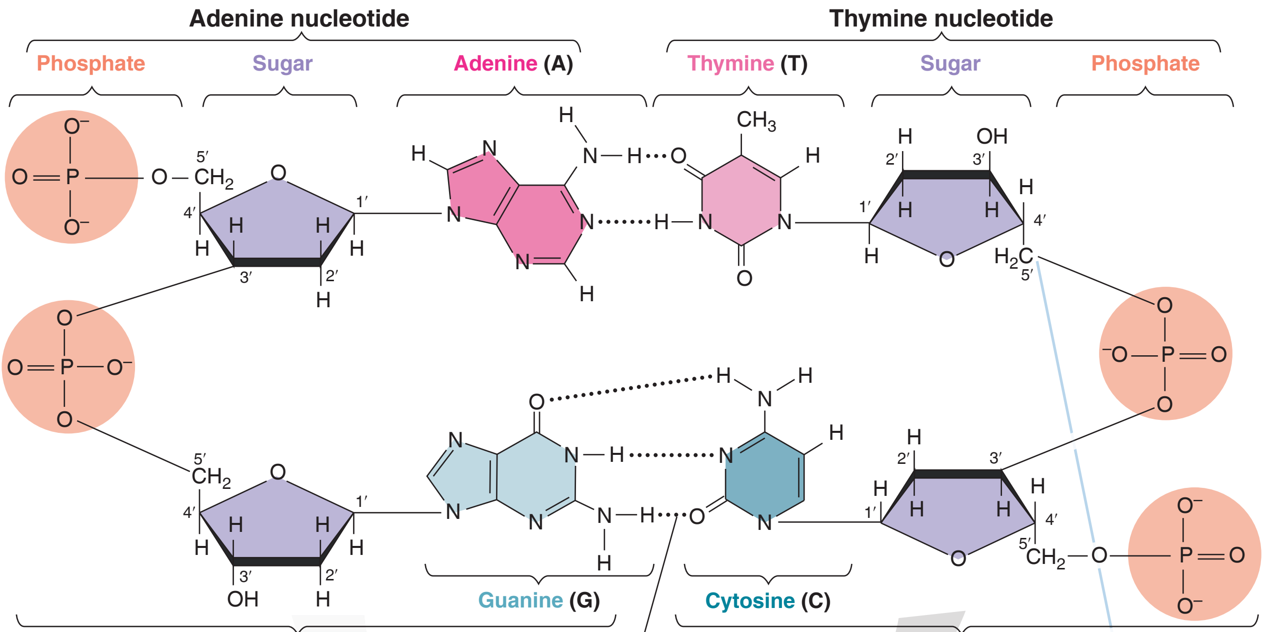
The order in which the nitrogen base pairs occur along the backbone is extremely specific and in fact contains the genetic instructions for the organism. Nucleotides form genes, and a single DNA molecule may contain thousands of genes. Genes determine all hereditary traits, and they control all the activities that take place within cells.

One very important consequence of nitrogen-containing base pairing is that if the sequence of bases of one strand is known, then the sequence of the other strand is also known. For example, if one strand has the sequence . . . ATGC . . . , then the other strand has the sequence . . . TACG . . . . Because the sequence of

FOUNDATION FIGURE 2.16

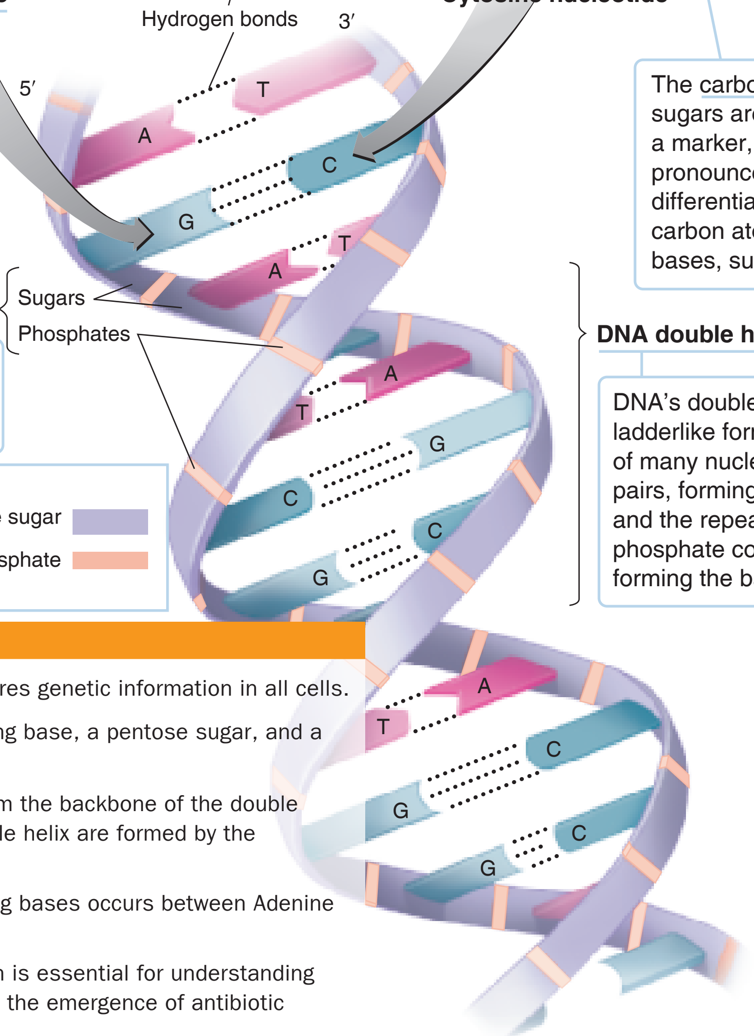
The Structure of DNA

Adenine, Thymine, Cytosine, and Guanine are nitrogenous bases, or nucleobases.



Individual DNA nucleotides are composed of a deoxyribose sugar molecule covalently bonded to a phosphate group at the 5′ carbon, and to a nitrogen-containing base at the 1′ carbon. The two nucleotides shown here are held together by hydrogen bonds.

The carbon atoms in the sugars are identified by adding a marker, ′ (for example, 5′, pronounced “5-prime”). This differentiates them from the carbon atoms in the nucleobases, such as Thymine.



Sugar-phosphate backbone

The sugar-phosphate backbone of one strand is upside down, or antiparallel, relative to the backbone of the other strand.

DNA double helix

DNA’s double-helical, ladderlike form is made up of many nucleotide base pairs, forming the rungs; and the repeating sugar-phosphate combination, forming the backbone.

Key

Adenine A — T Thymine
Guanine G — C Cytosine
Hydrogen bond ••••••••••
Deoxyribose sugar
Phosphate

KEY CONCEPTS

- DNA is a double-stranded molecule that stores genetic information in all cells.
- A nucleotide consists of a nitrogen-containing base, a pentose sugar, and a phosphate group.
- Alternating sugar and phosphate groups form the backbone of the double helix (twisted ladder); the rungs of the double helix are formed by the nitrogen-containing bases.
- Complementary pairing of nitrogen-containing bases occurs between Adenine and Thymine; Guanine and Cytosine.
- Familiarity with DNA’s structure and function is essential for understanding genetics, recombinant DNA techniques, and the emergence of antibiotic resistance and new diseases.

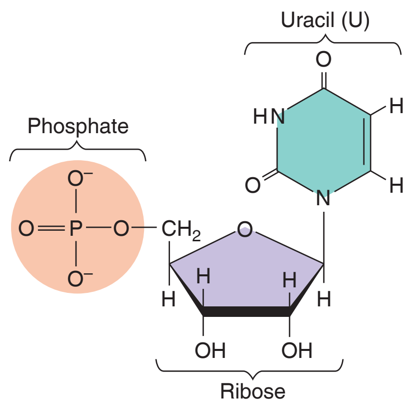


Figure 2.17 A uracil nucleotide of RNA.

Q How are DNA and RNA similar in structure?

bases of one strand is determined by the sequence of bases of the other, the bases are said to be *complementary*. The actual transfer of information becomes possible because of DNA's unique structure and will be discussed further in Chapter 8.

RNA

RNA, the second principal kind of nucleic acid, differs from DNA in several respects. Whereas DNA is double-stranded, RNA is usually single-stranded. The five-carbon sugar in the RNA nucleotide is ribose, which has one more oxygen atom than deoxyribose. Also, one of RNA's bases is uracil (U) instead of thymine (Figure 2.17). The other three bases (A, G, C) are the same as DNA. Three major kinds of RNA have been identified in cells. They are **messenger RNA (mRNA), ribosomal RNA (rRNA)**, and **transfer RNA (tRNA)**, each of which has a specific role in protein synthesis (see Chapter 8).

DNA and RNA are compared in Table 2.6.

CHECK YOUR UNDERSTANDING

✓ 2-11 How do DNA and RNA differ?

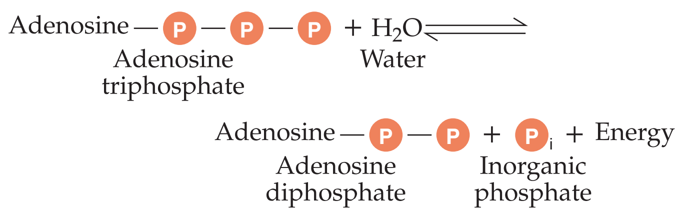
Adenosine Triphosphate (ATP)

Adenosine triphosphate (ATP) is the principal energy-carrying molecule of all cells and is indispensable to the life of the cell. It stores the chemical energy released by some chemical reactions, and it provides the energy for reactions that require energy. ATP consists of an adenosine unit, composed of adenine and ribose, with three phosphate groups (P) attached (Figure 2.18). In other words, it is an adenine nucleotide (also called adenosine monophosphate, or AMP) with two extra phosphate groups. ATP is called a high-energy molecule because it releases a large amount of usable energy when the third phosphate group is hydrolyzed to become **adenosine diphosphate (ADP)**. This reaction can be represented as follows:

CLINICAL CASE Resolved

The functional group in dipicolinic acid is carboxyl. *B. anthracis* infection is contracted by contact, ingestion, or inhalation of the endospores. In Jonathan's case, the process of stretching, scraping, and sanding the goat hides had created dust that settled on the drum skin and any surrounding crevices. *B. anthracis* endospores became airborne, or aerosolized, whenever Jonathan beat on the drum. He makes a full recovery, and from now on he makes certain that all parts of any drum he purchases have been legally imported.

25 42 44 **46**



A cell's supply of ATP at any particular time is limited. Whenever the supply needs replenishing, the reaction goes in the reverse direction; the addition of a phosphate group to ADP and the input of energy produces more ATP. The energy required to attach the terminal phosphate group to ADP is supplied by the cell's various oxidation reactions, particularly the oxidation of glucose. ATP can be made in every cell, where its potential energy is released when needed.

CHECK YOUR UNDERSTANDING

✓ 2-12 Which can provide more energy for a cell and why: ATP or ADP?

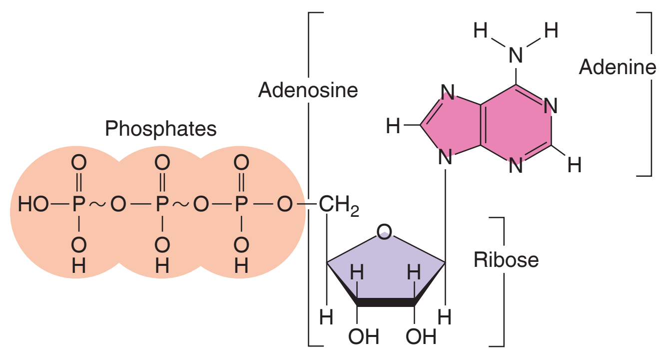


Figure 2.18 The structure of ATP. High-energy phosphate bonds are indicated by wavy lines. When ATP breaks down to ADP and inorganic phosphate, a large amount of chemical energy is released for use in other chemical reactions.

Q How is ATP similar to a nucleotide in RNA? In DNA?

TABLE 2.6 Comparison between DNA and RNA

Backbone	DNA	RNA
Strands	Double-stranded in cells and most DNA viruses to form a double helix; single-stranded in some viruses (parvoviruses).	Single-stranded in cells and most RNA viruses; double-stranded in some viruses (reoviruses).
Composition	The sugar is deoxyribose.	The sugar is ribose.
	The nitrogen-containing bases are cytosine (C), guanine (G), adenine (A), and thymine (T).	The nitrogen-containing bases are cytosine (C), guanine (G), adenine (A), and uracil (U).
Function	Determines all hereditary traits.	Protein synthesis.

Study Outline

Go to @MasteringMicrobiology for **Interactive Microbiology**, *In the Clinic videos, MicroFlix, MicroBoosters, 3D animations, practice quizzes, and more.*

Introduction (p. 24)

1. The science of the interaction between atoms and molecules is called chemistry.
2. The metabolic activities of microorganisms involve complex chemical reactions.
3. Microbes break down nutrients to obtain energy and to make new cells.

The Structure of Atoms (pp. 25–27)

1. An atom is the smallest unit of a chemical element that exhibits the properties of that element.
2. Atoms consist of a nucleus, which contains protons and neutrons, and electrons, which move around the nucleus.
3. The atomic number is the number of protons in the nucleus; the total number of protons and neutrons is the atomic mass.

Chemical Elements (p. 25)

4. Atoms with the same number of protons and the same chemical behavior are classified as the same chemical element.
5. Chemical elements are designated by abbreviations called chemical symbols.
6. About 26 elements are commonly found in living cells.
7. Atoms that have the same atomic number (are of the same element) but different atomic masses are called isotopes.

Electronic Configurations (p. 26)

8. In an atom, electrons are arranged around the nucleus in electron shells.
9. Each shell can hold a characteristic maximum number of electrons.
10. The chemical properties of an atom are due largely to the number of electrons in its outermost shell.

How Atoms Form Molecules: Chemical Bonds (pp. 27–29)

1. Molecules are made up of two or more atoms; molecules consisting of at least two different kinds of atoms are called compounds.
2. Atoms form molecules in order to fill their outermost electron shells.
3. Attractive forces that bind two atoms together are called chemical bonds.
4. The combining capacity of an atom—the number of chemical bonds the atom can form with other atoms—is its valence.

Ionic Bonds (p. 27)

5. A positively or negatively charged atom or group of atoms is called an ion.
6. A chemical attraction between ions of opposite charge is called an ionic bond.
7. To form an ionic bond, one ion is an electron donor, and the other ion is an electron acceptor.

Covalent Bonds (pp. 27–28)

8. In a covalent bond, atoms share pairs of electrons.
9. Covalent bonds are stronger than ionic bonds and are far more common in organic molecules.

Hydrogen Bonds (pp. 28–29)

10. A hydrogen bond exists when a hydrogen atom covalently bonded to one oxygen or nitrogen atom is attracted to another oxygen or nitrogen atom.
11. Hydrogen bonds form weak links between different molecules or between parts of the same large molecule.

Molecular Mass and Moles (p. 29)

12. The molecular mass is the sum of the atomic masses of all the atoms in a molecule.
13. A mole of an atom, ion, or molecule is equal to its atomic or molecular mass expressed in grams.

Chemical Reactions (pp. 30–31)

1. Chemical reactions are the making or breaking of chemical bonds between atoms.
2. A change of energy occurs during chemical reactions.
3. Endergonic reactions require more energy than they release; exergonic reactions release more energy.
4. In a synthesis reaction, atoms, ions, or molecules are combined to form a larger molecule.
5. In a decomposition reaction, a larger molecule is broken down into its component molecules, ions, or atoms.
6. In an exchange reaction, two molecules are decomposed, and their subunits are used to synthesize two new molecules.
7. The products of reversible reactions can readily revert to form the original reactants.

Important Biological Molecules (pp. 31–47)

Inorganic Compounds (pp. 31–33)

1. Inorganic compounds are usually small, ionically bonded molecules.

Water (pp. 31–32)

2. Water is the most abundant substance in cells.
3. Because water is a polar molecule, it is an excellent solvent.
4. Water is a reactant in many of the decomposition reactions of digestion.
5. Water is an excellent temperature buffer.

Acids, Bases, and Salts (p. 32)

6. An acid dissociates into H^+ and anions.
7. A base dissociates into OH^- and cations.
8. A salt dissociates into negative and positive ions, neither of which is H^+ or OH^-.

Acid–Base Balance: The Concept of pH (pp. 32–33)

9. The term *pH* refers to the concentration of H^+ in a solution.
10. A solution of pH 7 is neutral; a pH value below 7 indicates acidity; pH above 7 indicates alkalinity.
11. The pH inside a cell and in culture media is stabilized with pH buffers.

Organic Compounds (pp. 33–47)

1. Organic compounds always contain carbon and hydrogen.
2. Carbon atoms form up to four bonds with other atoms.
3. Organic compounds are mostly or entirely covalently bonded.

Structure and Chemistry (pp. 34–35)

4. A chain of carbon atoms forms a carbon skeleton.
5. Functional groups of atoms are responsible for most of the properties of organic molecules.
6. The letter *R* may be used to denote the remainder of an organic molecule.
7. Frequently encountered classes of molecules are R—OH (alcohols) and R—COOH (organic acids).
8. Small organic molecules may combine into very large molecules called macromolecules.
9. Monomers usually bond together by dehydration synthesis, or condensation reactions, that form water and a polymer.
10. Organic molecules may be broken down by hydrolysis, a reaction involving the splitting of water molecules.

Carbohydrates (pp. 35–36)

11. Carbohydrates are compounds consisting of atoms of carbon, hydrogen, and oxygen, with hydrogen and oxygen in a 2:1 ratio.
12. Monosaccharides contain from three to seven carbon atoms.
13. Isomers are two molecules with the same chemical formula but different structures and properties—for example, glucose ($C_6H_{12}O_6$) and fructose ($C_6H_{12}O_6$).
14. Monosaccharides may form disaccharides and polysaccharides by dehydration synthesis.

Lipids (pp. 36–40)

15. Lipids are a diverse group of compounds distinguished by their insolubility in water.
16. Simple lipids (fats) consist of a molecule of glycerol and three molecules of fatty acids.
17. A saturated lipid has no double bonds between carbon atoms in the fatty acids; an unsaturated lipid has one or more double bonds. Saturated lipids have higher melting points than unsaturated lipids.
18. Phospholipids are complex lipids consisting of glycerol, two fatty acids, and a phosphate group.
19. Steroids have carbon ring structures; sterols have a functional hydroxyl group.

Proteins (pp. 40–44)

20. Amino acids are the building blocks of proteins.
21. Amino acids consist of carbon, hydrogen, oxygen, nitrogen, and sometimes sulfur.
22. Twenty amino acids occur naturally in proteins.
23. By linking amino acids, peptide bonds (formed by dehydration synthesis) allow the formation of polypeptide chains.
24. Proteins have four levels of structure: primary (sequence of amino acids), secondary (helices or pleated), tertiary (overall three-dimensional structure of a polypeptide), and quaternary (two or more polypeptide chains).
25. Conjugated proteins consist of amino acids combined with inorganic or other organic compounds.

Nucleic Acids (pp. 44–46)

26. Nucleic acids—DNA and RNA—are macromolecules consisting of repeating nucleotides.
27. A nucleotide is composed of a pentose, a phosphate group, and a nitrogen-containing base. A nucleoside is composed of a pentose and a nitrogen-containing base.
28. A DNA nucleotide consists of deoxyribose (a pentose) and one of the following nitrogen-containing bases: thymine or cytosine (pyrimidines) or adenine or guanine (purines).
29. DNA consists of two strands of nucleotides wound in a double helix. The strands are held together by hydrogen bonds between purine and pyrimidine nucleotides: AT and GC.
30. Genes consist of sequences of nucleotides.
31. An RNA nucleotide consists of ribose (a pentose) and one of the following nitrogen-containing bases: cytosine, guanine, adenine, or uracil.

Adenosine Triphosphate (ATP) (p. 46)

32. ATP stores chemical energy for various cellular activities.
33. When the bond to ATP's terminal phosphate group is hydrolyzed, energy is released.
34. The energy from oxidation reactions is used to regenerate ATP from ADP and inorganic phosphate.

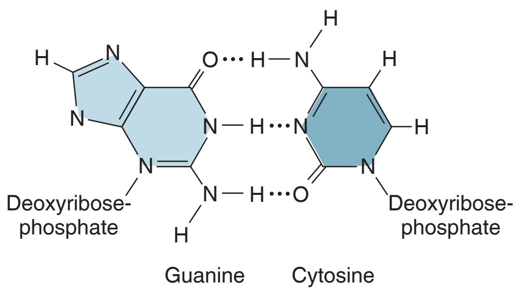
Study Questions

For answers to the Knowledge and Comprehension questions, turn to the Answers tab at the back of the textbook.

Knowledge and Comprehension

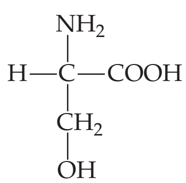
Review

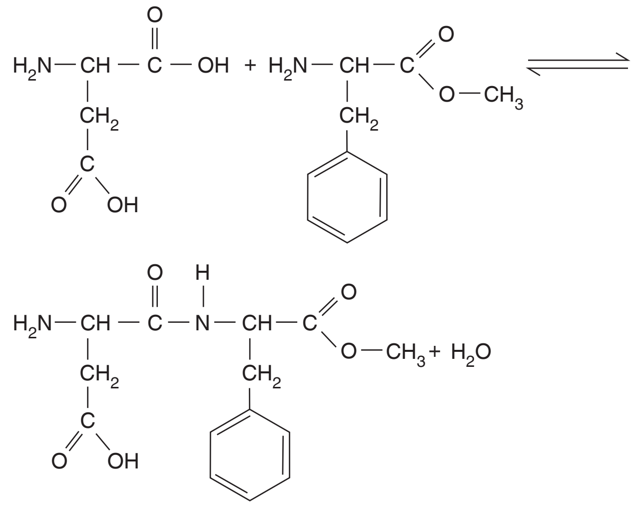
1. What is a chemical element?
2. DRAW IT Diagram the electronic configuration of a carbon atom.
3. What type of bond holds the following atoms together?
   a. Li^+ and Cl^- in LiCl
   b. carbon and oxygen atoms in methanol
   c. oxygen atoms in O_2
   d. a hydrogen atom of one nucleotide to a nitrogen or oxygen atom of another nucleotide in:



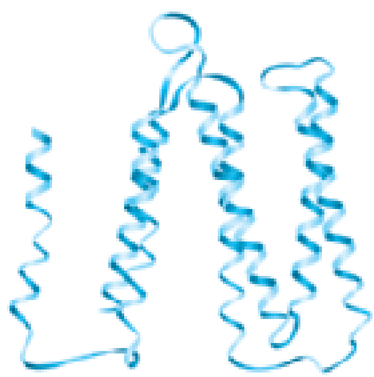
4. Classify the following types of chemical reactions.
   a. glucose + fructose → sucrose + H_2O
   b. lactose → glucose + galactose
   c. $NH_4Cl + H_2O \rightarrow NH_4OH + HCl$
   d. ATP ⇌ ADP+ P_i
5. Bacteria use the enzyme urease to obtain nitrogen in a form they can use from urea in the following reaction:

$$\underset{\text{Urea}}{CO(NH_2)_2} + H_2O \rightarrow \underset{\text{Ammonia}}{2NH_3} + \underset{\text{Carbon dioxide}}{CO_2}$$

   What purpose does the enzyme serve in this reaction? What type of reaction is this?
6. Classify the following as subunits of either a carbohydrate, lipid, protein, or nucleic acid.
   a. $CH3—(CH_2)_7—CH=CH—(CH_2)_7—COOH$
      Oleic acid
   b. 
      Serine
   c. $C_6H_{12}O_6$
   d. Thymine nucleotide
7. DRAW IT The artificial sweetener aspartame, or NutraSweet®, is made by joining aspartic acid to methylated phenylalanine, as shown below.



   a. What types of molecules are aspartic acid and phenylalanine?
   b. What direction is the hydrolysis reaction (left to right or right to left)?
   c. What direction is the dehydration synthesis reaction?
   d. Circle the atoms involved in the formation of water.
   e. Identify the peptide bond.
8. DRAW IT The following diagram shows the bacteriorhodopsin protein. Indicate the regions of primary, secondary, and tertiary structure. Does this protein have quaternary structure?



9. DRAW IT Draw a simple lipid, and show how it could be modified to a phospholipid.
10. NAME IT What type of microorganism has a chitin cell wall, has DNA that is contained in a nucleus, and has ergosterol in its plasma membrane?

Multiple Choice

Radioisotopes are frequently used to label molecules in a cell. The fate of atoms and molecules in a cell can then be followed. This process is the basis for questions 1–3.

1. Assume *E. coli* bacteria are grown in a nutrient medium containing the radioisotope ^{16}N. After a 48-hour incubation period, the ^{16}N would most likely be found in the *E. coli*'s
   a. carbohydrates.
   b. lipids.
   c. proteins.
   d. water.
   e. none of the above
2. If *Pseudomonas* bacteria are supplied with radioactively labeled cytosine, after a 24-hour incubation period this cytosine would most likely be found in the cells'
   a. carbohydrates.
   b. DNA.
   c. lipids.
   d. water.
   e. proteins.
3. If *E. coli* were grown in a medium containing the radioactive isotope ^{32}P, the ^{32}P would be found in all of the following molecules of the cell *except*
   a. ATP.
   b. carbohydrates.
   c. DNA.
   d. plasma membrane.
   e. none of the above
4. The optimum pH of *Acidithiobacillus* bacteria (pH 3,) is _________ times more acid than blood (pH 7).
   a. 4
   b. 10
   c. 100
   d. 1000
   e. 10,000
5. The best definition of ATP is that it is
   a. a molecule stored for food use.
   b. a molecule that supplies energy to do work.
   c. a molecule stored for an energy reserve.
   d. a molecule used as a source of phosphate.
6. Which of the following is an organic molecule?
   a. H_2O (water)
   b. O_2 (oxygen)
   c. $C_{18}H_{29}SO_3$ (Styrofoam)
   d. FeO (iron oxide)
   e. $F_2C{=}CF_2$ (Teflon)

Classify each of the molecules on the left as an acid, base, or salt. The dissociation products of the molecules are shown to help you.

7. $HNO_3 \rightarrow H^+ + NO_3^-$
8. $H_2SO_4 \rightarrow 2H^+ + SO_4^{2-}$
9. $NaOH \rightarrow Na^+ + OH^-$
10. $MgSO_4 \rightarrow Mg^{2+} + SO_4^{2-}$

a. acid
b. base
c. salt

Analysis

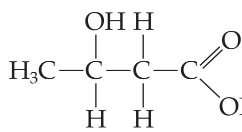
1. When you blow bubbles into a glass of water, the following reactions take place:

$$H_2O + CO_2 \xrightarrow{A} H_2CO_3 \xrightarrow{B} H^+ + HCO_3^-$$

   a. What type of reaction is *A*?
   b. What does reaction *B* tell you about the type of molecule H_2CO_3 is?
2. What are the common structural characteristics of ATP and DNA molecules?
3. What happens to the relative amount of unsaturated lipids in the plasma membrane when *E. coli* bacteria grown at 25°C are then grown at 37°C?
4. Giraffes, termites, and koalas eat only plant matter. Because animals cannot digest cellulose, how do you suppose these animals get nutrition from the leaves and wood they eat?

Clinical Applications and Evaluation

1. *Ralstonia* bacteria make poly-β-hydroxybutyrate (PHB), which is used to make a biodegradable plastic. PHB consists of many of the monomers shown below. What type of molecule is PHB? What is the most likely reason a cell would store this molecule?



2. *Acidithiobacillus ferrooxidans* was responsible for destroying buildings in the Midwest by causing changes in the earth. The original rock, which contained lime ($CaCO_3$) and pyrite (FeS_2), expanded as bacterial metabolism caused gypsum ($CaSO_4$) crystals to form. How did *A. ferrooxidans* bring about the change from lime to gypsum?
3. Newborn babies are tested for phenylketonuria (PKU), an inherited disease. Individuals with this disease are missing an enzyme to convert phenylalanine (phe) to tyrosine; the resulting accumulation of phe can cause mental retardation, brain damage, and seizures. The Guthrie test for PKU involves culturing *Bacillus subtilis*, which requires phe to grow. The bacteria are grown on media with a drop of the baby's blood.
   a. What type of chemical is phenylalanine?
   b. What does "no growth" in the Guthrie test mean?
   c. Why must individuals with PKU avoid the sweetener aspartame?
4. The antibiotic amphotericin B causes leaks into cells by combining with sterols in the plasma membrane. Would you expect to use amphotericin B against a bacterial infection? A fungal infection? Offer a reason why amphotericin B has severe side effects in humans.
5. You can smell sulfur when boiling eggs. What amino acids do you expect in the egg?