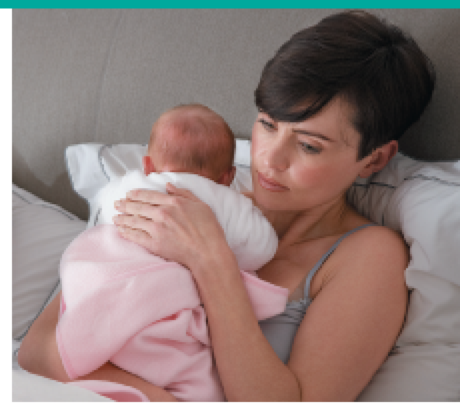
2 The Chemical Level of Organization

Learning Outcomes

These Learning Outcomes correspond by number to this chapter's sections and indicate what you should be able to do after completing the chapter.

CLINICAL CASE What Is Wrong with My Baby?



Sean is Maureen's first baby. Maureen and her husband, Conner, had enjoyed an uncomplicated pregnancy and delivery. Maureen had felt healthy throughout the pregnancy, but something was wrong with her baby.

Sean is 1 month old, and not thriving. He seems to have a good appetite and breast-feeds as if he were starving; yet he has dropped 20 percent of his normal birth weight of 7 pounds, 8 ounces. He is down to only 6 pounds at his 1-month checkup, and his skin looks "wrinkly." His stools appear greasy and foamy. Maureen also notices that Sean's skin tastes salty.

Most alarmingly, he seems to be having some difficulty breathing. His breathing is wheezy. Maureen and Conner are both from big families, and none of the babies has ever been sickly like this. **What is wrong with baby Sean? To find out, turn to the Clinical Case Wrap-Up on p. 64.**

An Introduction to the Chemical Level of Organization

In this chapter we consider the structure of *atoms*, the basic chemical building blocks. You will also learn how atoms can combine to form increasingly complex structures, and how those types of complex structures function in the human body.

2-1 Atoms are the basic particles of matter

Learning Outcome Describe an atom and how atomic structure affects interactions between atoms.

Our study of the human body begins at the chemical level of organization. **Chemistry** is the science that deals with the structure of **matter**, defined as anything that takes up space and has mass. **Mass**, the amount of material in matter, is a physical property that determines the weight of an object in Earth's gravitational field. For our purposes, the mass of an object is the same as its weight. However, the two are not always equivalent: In orbit you would be weightless, but your mass would remain unchanged.

The smallest stable units of matter are called **atoms**. Air, elephants, oranges, oceans, rocks, and people are all composed of atoms in varying combinations. The unique characteristics of each object, living or nonliving, result from the types of atoms involved and the ways those atoms combine and interact.

Atoms are composed of **subatomic particles**. Many different subatomic particles exist, but only three—*protons, neutrons*, and *electrons*—are important for understanding the chemical properties of matter. Protons and neutrons are similar in size and mass, but **protons** (p^+) have a positive electrical charge. **Neutrons** (n or n^0) are electrically *neutral*, or uncharged. **Electrons** (e^-) are much lighter than protons—only 1/1836 as massive—and have a negative electrical charge. For this reason, the mass of an atom is determined primarily by the number of protons and neutrons in the **nucleus**, the central region of an atom. The mass of a large object, such as your body, is the sum of the masses of all its component atoms.

Atomic Structure

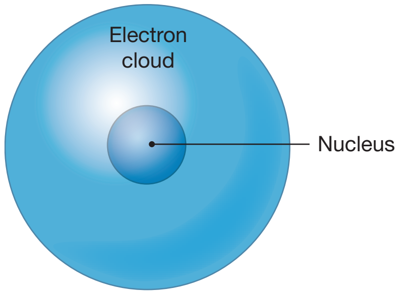
Atoms normally contain equal numbers of protons and electrons. The number of protons in an atom is known as its **atomic number**. *Hydrogen* (represented as H) is the simplest atom, with an atomic number of 1. An atom of hydrogen contains one proton and one electron. Hydrogen's proton is located in the center of the atom and forms the nucleus. Hydrogen atoms seldom contain neutrons, but when neutrons are present, they are also located in the nucleus. All atoms other than hydrogen have both neutrons and protons in their nuclei.

Electrons travel around the nucleus at high speed, within a spherical area called the **electron cloud** (Figure 2–1). We often illustrate atomic structure in the simplified form shown for hydrogen in Figure 2–2a (see p. 30). In this two-dimensional representation, the electrons occupy a circular **electron shell**. One reason an electron tends to remain in its electron shell is that the negatively charged electron is attracted to the positively charged proton. The attraction between opposite electrical charges is an example of an *electrical force*. As you will see in later chapters, electrical forces are involved in many physiological processes.

The dimensions of the electron cloud determine the overall size of the atom. To get an idea of the scale involved, consider that

> Go to MasteringA&P™ > Study Area > Get Ready for A&P > 2. Get Ready for A&P Website> Chapter 5: Chemistry > Atomic Structure: Video Tutor: Atomic Structure

Figure 2–1 Hydrogen Atom with Electron Cloud. This space-filling model of a hydrogen atom depicts the three-dimensional electron cloud formed by the single electron orbiting the nucleus.



if the nucleus were the size of a tennis ball, the electron cloud of a hydrogen atom would have a radius of 10 km (about 6 miles!). In reality, atoms are so small that atomic measurements are reported in nanometers (NAN-ō-mē-terz) (nm). One nanometer is 10^{-9} meter (0.000000001 m), or one billionth of a meter. The very largest atoms approach 0.5 nm in diameter.

Elements and Isotopes

An **element** is a pure substance composed of atoms of only one kind. Atoms are the smallest particles of an element that still retain the characteristics of that element. As a result, each element has uniform composition and properties. Each element includes all the atoms with the same number of protons, and thus the same atomic number. Only 92 elements exist in nature. Researchers have created about two dozen additional elements through nuclear reactions in laboratories.

Every element has a chemical symbol, an abbreviation recognized by scientists everywhere. Most of the symbols are easy to connect with the English names of the elements (O for oxygen, N for nitrogen, C for carbon, and so on), but a few are abbreviations of their Latin names. For example, the symbol for sodium, Na, comes from the Latin word *natrium*.

Elements cannot be changed or broken down into simpler substances, whether by chemical processes, heating, or other ordinary physical means. For example, an atom of carbon always remains an atom of carbon, regardless of the chemical events in which it may take part.

Our bodies consist of many elements, and the 13 most abundant elements are listed in **Table 2–1**. Our bodies also contain atoms of another 14 elements—called *trace elements*—that are present in very small amounts.

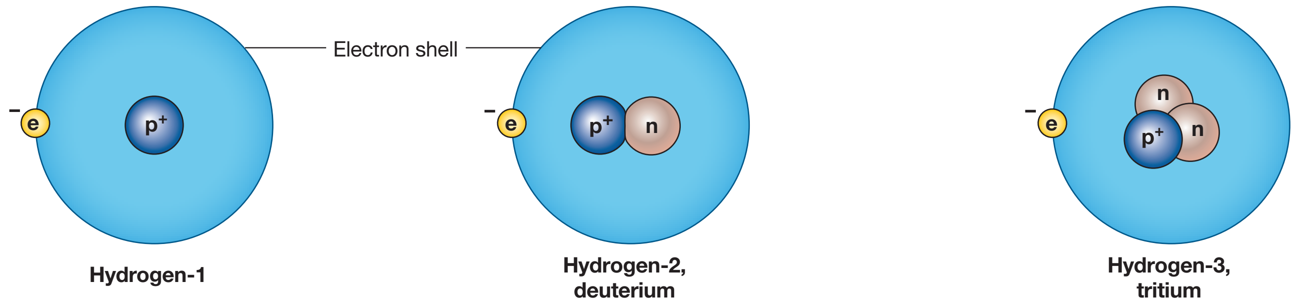
The atoms of a single element all have the same number of protons, but they can differ in the number of neutrons in the nucleus. For example, most hydrogen nuclei consist of just a single proton, but 0.015 percent also contain 1 neutron, and a very small percentage contain 2 neutrons (**Figure 2–2**). Atoms of the same element whose nuclei contain different numbers of neutrons are called **isotopes**.

Different isotopes of an element have essentially identical chemical properties, and are alike except in mass. The **mass number**—the total number of protons plus neutrons in the nucleus of an atom—is used to designate isotopes. For example, hydrogen has 3 isotopes, distinguished by their mass numbers (1, 2, or 3). Hydrogen-1, or ^{1}H, has 1 proton and 1 electron (**Figure 2–2a**). Hydrogen-2, or ^{2}H, also known as *deuterium*, has 1 proton, 1 electron, and 1 neutron (**Figure 2–2b**). Hydrogen-3,

Table 2–1 Principal Elements in the Human Body

Element (% of total body weight)	Significance
Oxygen, O (65)	A component of water and other compounds; gaseous form is essential for respiration
Carbon, C (18.6)	Found in all organic molecules
Hydrogen, H (9.7)	A component of water and most other compounds in the body
Nitrogen, N (3.2)	Found in proteins, nucleic acids, and other organic compounds
Calcium, Ca (1.8)	Found in bones and teeth; important for membrane function, nerve impulses, muscle contraction, and blood clotting
Phosphorus, P (1.0)	Found in bones and teeth, nucleic acids, and high-energy compounds
Potassium, K (0.4)	Important for proper membrane function, nerve impulses, and muscle contraction
Sodium, Na (0.2)	Important for blood volume, membrane function, nerve impulses, and muscle contraction
Chlorine, Cl (0.2)	Important for blood volume, membrane function, and water absorption
Magnesium, Mg (0.06)	A cofactor for many enzymes
Sulfur, S (0.04)	Found in many proteins
Iron, Fe (0.007)	Essential for oxygen transport and energy capture
Iodine, I (0.0002)	A component of hormones of the thyroid gland
Trace elements: silicon (Si), fluorine (F), copper (Cu), manganese (Mn), zinc (Zn), selenium (Se), cobalt (Co), molybdenum (Mo), cadmium (Cd), chromium (Cr), tin (Sn), aluminum (Al), boron (B), and vanadium (V)	Some function as cofactors; the functions of many trace elements are poorly understood

Figure 2–2 The Structure of Hydrogen Atoms. Three forms of hydrogen atoms are shown using the two-dimensional electron-shell model, which indicates the spherical electron cloud surrounding the nucleus.



a A typical hydrogen nucleus contains 1 proton and no neutrons.

b A deuterium (^2H) nucleus contains 1 proton and 1 neutron.

c A tritium (^3H) nucleus contains 1 proton and 2 neutrons.

or ^3H, also known as *tritium*, has 1 proton, 1 electron, and 2 neutrons (**Figure 2–2c**).

The nuclei of some isotopes are unstable, or radioactive. That is, they spontaneously break down and give off *radiation* (energy in the form of moving subatomic particles or waves) in measurable amounts. Such isotopes are called **radioisotopes**. The breakdown process is called *radioactive decay*. The decay rate of a radioisotope is commonly expressed as its **half-life**: the time required for half of a given amount of the isotope to decay. Radioisotopes differ radically in how rapidly they decay; their half-lives range from fractions of a second to billions of years.

Weakly radioactive isotopes are sometimes used in diagnostic procedures to monitor the structural or functional characteristics of internal organs. However, strongly radioactive isotopes are dangerous, because the radiation they give off can alter the number of electrons in an atom, break apart molecules, and destroy cells and tissues.

Atomic Weights

A typical *oxygen* atom has an atomic number of 8 and contains 8 protons and 8 neutrons. The mass number of this isotope is therefore 16. The mass numbers of other isotopes of oxygen depend on the number of neutrons present. Mass numbers are useful because they tell us the number of subatomic particles in the nuclei of different atoms. However, they do not tell us the *actual* mass of the atoms. For example, they do not take into account the masses of the electrons or the slight difference between the mass of a proton and that of a neutron. The actual mass of an atom of a specific isotope is known as its *atomic mass*.

The unit used to express atomic mass is the *atomic mass unit* (amu), or *dalton*. By international agreement, 1 amu is equal to one-twelfth the mass of a carbon-12 atom. One atomic mass unit is very close to the mass of a single proton or neutron. Thus, the atomic mass of an atom of the most common isotope of hydrogen is very close to 1, and that of the most common isotope of oxygen is very close to 16.

The **atomic weight** of an element is an average of the different atomic masses and proportions of its different isotopes. This results in the atomic weight of an element being very close to the mass number of the most common isotope of that element. For example, the mass number of the most common isotope of hydrogen is 1, but the atomic weight of hydrogen is closer to 1.01, primarily because some hydrogen atoms (0.02 percent) have a mass number of 2, and even fewer have a mass number of 3. (The periodic table of the elements in the Appendix at the back of this book shows the atomic weight of each element.)

Atoms take part in chemical reactions in fixed numerical ratios. To form water, for example, exactly 2 atoms of hydrogen combine with 1 atom of oxygen. But individual atoms are far too small and too numerous to be counted, so to determine the number of atoms chemists use a unit called the *mole*. For any element, a **mole** (abbreviated *mol*) is a specific quantity with a weight in grams equal to that element's atomic weight.

The mole is useful because 1 mol of a given element always contains the same number of atoms as 1 mol of any other element (just as we use *dozen* to stand for 12 items). That number (called *Avogadro's number*) is 6.023×10^{23}, or about 600 billion trillion. Expressing relationships in moles rather than in grams makes it much easier to keep track of the relative numbers of atoms in chemical samples and processes. For example, if a report stated that a sample contains 0.5 mol of hydrogen atoms and 0.5 mol of oxygen atoms, you would know immediately that the 2 elements were present in equal numbers. That would not be so evident if the report stated that there were 0.505 g of hydrogen atoms and 8.00 g of oxygen atoms. Most chemical analyses and clinical laboratory tests report data in moles (mol), millimoles (mmol—1/1000 mol, or 10^{-3} mol), or micromoles (μmol—1/1,000,000 mol, or 10^{-6} mol).

Electrons and Energy Levels

Atoms are electrically neutral. In other words, every positively charged proton is balanced by a negatively charged electron. Thus, each increase in the atomic number has a comparable increase in the number of electrons traveling around the nucleus.

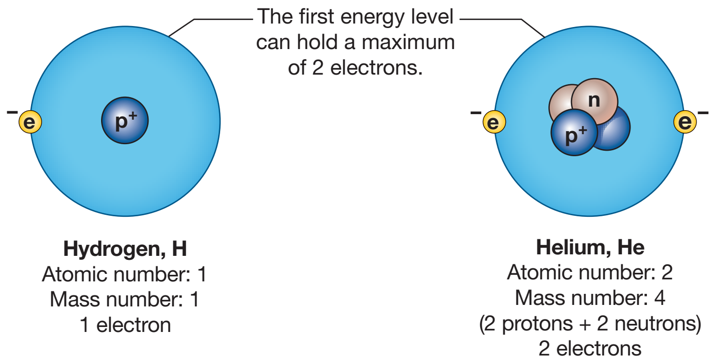
Within the electron cloud, electrons occupy an orderly series of *energy levels*. The electrons in an energy level may travel in complex patterns around the nucleus, but for our purposes the patterns can be diagrammed as a series of concentric electron shells. The first electron shell is the one closest to the nucleus, and it corresponds to the lowest energy level of electrons. Note that the terms *electron shell* and *energy level* can generally be used interchangeably.

There are up to eight energy levels in atoms, depending on their atomic number, but let's just look at the first three here. Each energy level is limited in the number of electrons it can hold. The first energy level can hold at most 2 electrons. The next two levels can each hold up to 8 electrons, an observation called the *octet rule*. Note that the maximum number of electrons that may occupy shells 1 through 3 corresponds to the number of elements in rows 1 through 3 of the periodic table of the elements (see Appendix). The electrons in an atom occupy successive shells in an orderly manner: The first energy level fills before any electrons enter the second, and the second energy level fills before any electrons enter the third.

The outermost energy level, or electron shell, forms the "surface" of the atom and is called the **valence shell**. (The *valence* of an element refers to its combining power with other atoms.) The number of electrons in this level determines the chemical properties of the element. Atoms with unfilled valence shells are unstable—that is, they will react with other atoms, usually in ways that result in full valence shells. In contrast, atoms with a filled valence shell are stable and therefore do not readily react with other atoms.

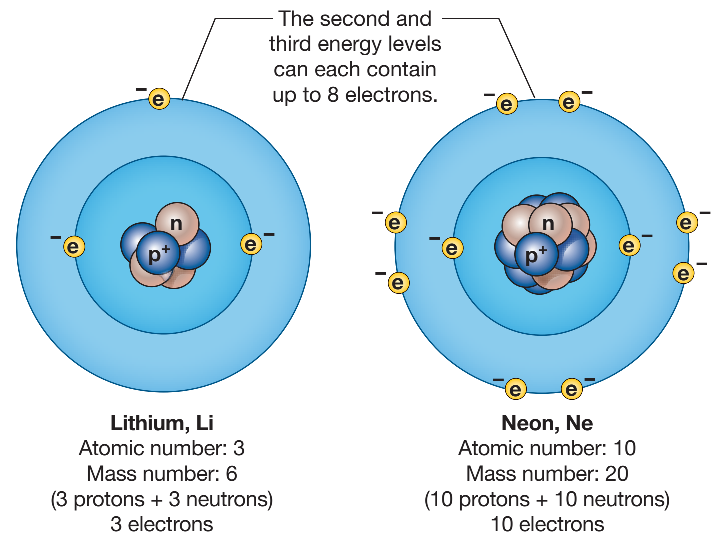
As indicated in **Figure 2–3a**, a hydrogen atom has 1 electron in the first energy level, the valence shell, so that level is unfilled. Therefore, a hydrogen atom readily reacts with other atoms. A helium atom has 2 electrons in its first energy level, which means its valence shell is filled (**Figure 2–3b**). This makes the helium atom very stable; it will not ordinarily react with other atoms. A lithium atom has 3 electrons (**Figure 2–3c**). Its first energy level can hold only 2 of them, so lithium has a single electron in a second, unfilled energy level. Thus, like hydrogen, lithium is unstable and reactive. The valence shell is filled in a neon atom, which has an atomic number of 10 (**Figure 2–3d**). Neon atoms, like helium atoms and other elements in the far right column of the periodic table, are very stable. The atoms of elements that are most important to biological systems are unstable (see **Table 2–1**, p. 29). Their instability promotes atomic interactions to form larger structures.

Figure 2–3 The Arrangement of Electrons into Energy Levels. The electron-shell model is also used to indicate the relative energy of electrons in an atom.



a **Hydrogen (H).** A typical hydrogen atom has 1 proton and 1 electron. The electron orbiting the nucleus occupies the first, or lowest, energy level, diagrammed as an electron shell.

b **Helium (He).** An atom of helium has 2 protons, 2 neutrons, and 2 electrons. The 2 electrons orbit in the same energy level.



c **Lithium (Li).** A lithium atom has 3 protons, 3 neutrons, and 3 electrons. The first energy level can hold only 2 electrons, so the third electron occupies the second energy level.

d **Neon (Ne).** A neon atom has 10 protons, 10 neutrons, and 10 electrons. The second level can hold up to 8 electrons; thus, both the first and second energy levels are filled.

? How many electrons can the second energy level hold when it is completely filled?

2

Clinical Note Radiation Sickness

Radiation sickness results from excessive exposure to ionizing radiation. It is characterized by fatigue, nausea, vomiting, and loss of teeth and hair. More severe cases cause anemia (low red blood cell count), central nervous damage, and death. In a clinical setting, it can occur when large doses of cancer-treating radiation are given to a person over a short time period. (X-rays and CT scans use low-dose radiation and do not cause radiation sickness.) The amount of radiation received determines how sick a person can become. In 2011, a tsunami damaged a nuclear reactor in Fukushima, Japan, releasing a lethal radioactive form of iodine, I-131. Since then, public health officials have made potassium iodide (KI) available to people who live near such reactors. This benign salt saturates the thyroid gland with stable (nonradioactive) iodine so that it cannot absorb I-131.

Checkpoint

1. **Define *atom*.**
2. **Atoms of the same element that have different numbers of neutrons are called _____.**
3. **How is it possible for two samples of hydrogen to contain the same number of atoms, yet have different weights?**

See the blue Answers tab at the back of the book.

2-2 Chemical bonds are forces formed by interactions between atoms

Learning Outcome Compare the ways in which atoms combine to form molecules and compounds.

Elements without active chemical properties are said to be *inert*. The noble gases helium, neon, and argon have filled valence shells and are called *inert gases*, because their atoms do not undergo chemical reactions.

Elements with unfilled valence shells, such as hydrogen and lithium, are called *reactive*, because they readily interact or combine with other atoms. Reactive atoms become stable by gaining, losing, or sharing electrons to fill their valence shells. The interactions often involve the formation of **chemical bonds**, which hold the participating atoms together once the chemical reaction has ended.

When chemical bonding takes place, new chemical entities called *molecules* and *compounds* are created. Chemical bonding may or may not involve the sharing of electrons. The term **molecule** refers to any chemical structure consisting of atoms held together by shared electrons. A **compound** is a pure chemical substance made up of atoms of two or more different elements in a fixed proportion, regardless of whether electrons are shared or not. The two categories overlap, but they are not the same. Not all molecules are compounds, because some molecules consist of atoms of only one element. (For example, two oxygen atoms can be joined by sharing electrons to form a molecule of oxygen.) And not all compounds consist of molecules, because some compounds, such as ordinary table salt (sodium chloride) are held together by bonds that do not involve shared electrons.

Many substances, however, fit both categories. Take water, for example. Water is a compound because it contains two different elements—hydrogen and oxygen—in a fixed proportion of two hydrogen atoms to one oxygen atom. It also consists of molecules, because the two hydrogen atoms and one oxygen atom are held together by shared electrons. As we will see in later sections, most biologically important compounds, from carbohydrates to DNA, are molecules.

Regardless of the type of bonding involved, a chemical compound has properties that can be quite different from those of its components. For example, a mixture of hydrogen gas and oxygen gas can explode, but the explosion is a chemical reaction that produces liquid water, a compound used to put out fires.

The **molecular weight** of a molecule or compound is the sum of the atomic weights of its component atoms. It follows from the definition of the mole given previously (p. 30) that the molecular weight of a molecule or compound in grams is equal to the mass of 1 mol of molecules or a compound. Molecular weights are important because we cannot handle individual molecules or parts of a compound nor can we easily count the billions involved in chemical reactions in the body.

As an example of how to calculate a molecular weight, let's use water, or H_2O. The atomic weight of hydrogen is 1.0079. To simplify our calculations, we round that value to 1. Then, 1 hydrogen molecule (H_2) with its 2 atoms has a molecular weight of $2(1 \times 2 = 2)$. One oxygen atom has an atomic weight of 16. Summing up, the molecular weight of 1 molecule of H_2O is 18. One mole of water molecules would then have a mass of 18 grams.

The human body consists of countless molecules and compounds, so it is a challenge to describe these substances and their varied interactions. Chemists simplify such descriptions through a standardized system of *chemical notation*. The basic rules of this system for atoms and molecules are described in **Spotlight Figure 2–4a,b**.

SPOTLIGHT

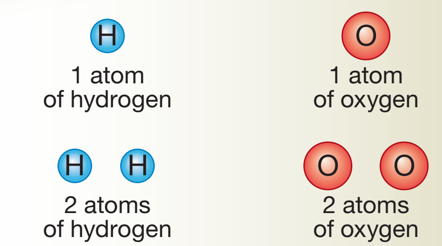
Figure 2–4
Chemical Notation

In order to discuss the specific compounds that occur in the human body, we must be able to describe chemical compounds and reactions clearly. To do this, we use a simple form of “chemical shorthand” known as **chemical notation**. Chemical notation enables us to describe complex events briefly and precisely; its rules are summarized below.

VISUAL REPRESENTATION | **CHEMICAL NOTATION**

a Atoms

The symbol of an element indicates 1 atom of that element. A number preceding the symbol of an element indicates more than 1 atom of that element.



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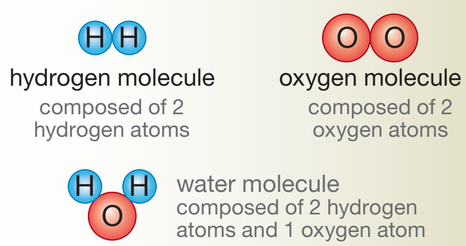
H — 1 atom of hydrogen; O — 1 atom of oxygen

2H — 2 atoms of hydrogen; 2O — 2 atoms of oxygen

b Molecules

A numerical subscript following the symbol of an element indicates the number of atoms of that element in a molecule.



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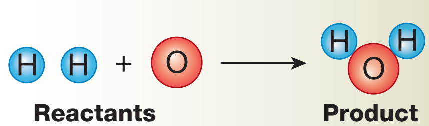
H_2 — hydrogen molecule; O_2 — oxygen molecule

H_2O — water molecule

c Reactions

In a description of a chemical reaction, the participants at the start of the reaction are called reactants, and the reaction generates one or more products. Chemical reactions are represented by chemical equations. An arrow indicates the direction of the reaction, from reactants (usually on the left) to products (usually on the right). In the following reaction, 2 atoms of hydrogen combine with 1 atom of oxygen to produce a single molecule of water.



Chemical reactions neither create nor destroy atoms; they merely rearrange atoms into new combinations. Therefore, the numbers of atoms of each element must always be the same on both sides of the equation for a chemical reaction. When this is the case, the equation is balanced.

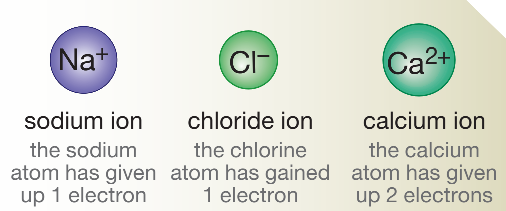
=

$2H + O \longrightarrow H_2O$
Balanced equation

$2H + 2O \longrightarrow H_2O$
Unbalanced equation

d Ions

A superscript plus or minus sign following the symbol of an element indicates an ion. A single plus sign indicates a cation with a charge of +1. (The original atom has given up 1 electron.) A single minus sign indicates an anion with a charge of –1. (The original atom has gained 1 electron.) If more than 1 electron is involved, the charge on the ion is indicated by a number preceding the plus or minus sign.



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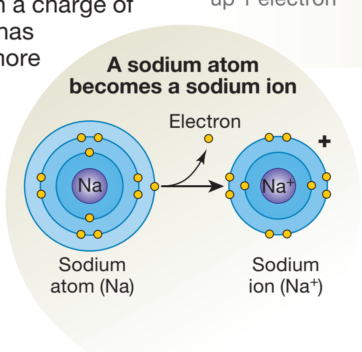
Na^+ — sodium ion; Cl^- — chloride ion; Ca^{2+} — calcium ion



In the sections that follow, we consider three basic types of chemical bonds: *ionic bonds*, *covalent bonds*, and *hydrogen bonds*.

Ionic Bonds

2

An **ion** is an atom or group of atoms that has an electric charge, either positive or negative. Ions with a positive charge (+) are called **cations** (KAT-ī-onz). Ions with a negative charge (−) are called **anions** (AN-ī-onz).

Tips & Tools

Think of the ***t*** in ca*t*ion as a plus sign (+) to remember that a ca*t*ion has a positive charge, and think of the ***n*** in a*n*ion as standing for ***n***egative (−) to remember that a*n*ions have a ***n***egative charge.

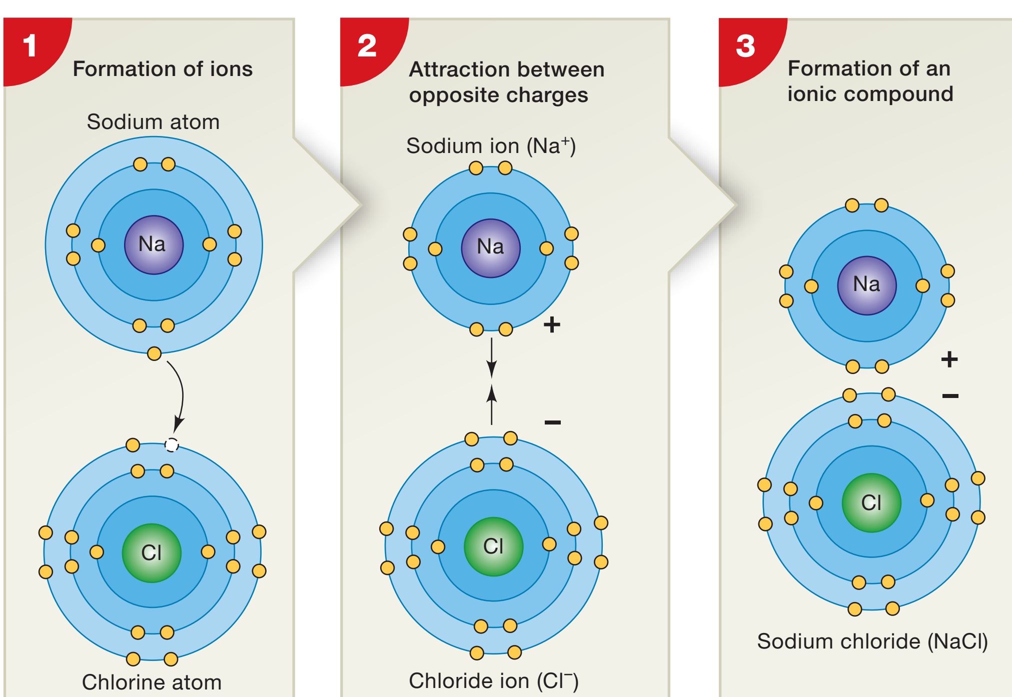
Atoms become ions by losing or gaining electrons, so ions have an unequal number of protons and electrons. We assign a value of +1 to the charge on a proton, while the charge on an electron is −1. An atom that loses an electron becomes a cation with a charge of +1, because it then has 1 proton that lacks a corresponding electron. Losing a second electron would give the cation a charge of +2. Adding an extra electron to a neutral atom produces an anion with a charge of −1. Adding a second electron gives the anion a charge of −2 (**Spotlight Figure 2–4d**).

Ionic bonds are chemical bonds created by the electrical attraction between anions and cations. In the formation of an ionic bond, the following steps occur:

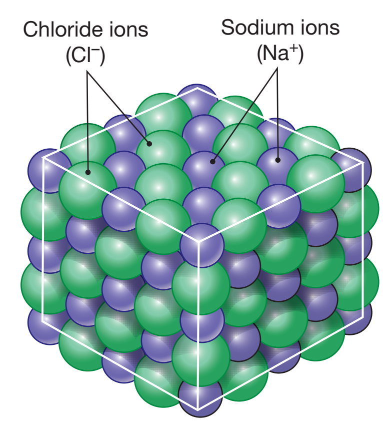
1. One atom—the *electron donor*—loses one or more electrons and becomes a cation, with a positive (+) charge. Another atom—the *electron acceptor*—gains those same electrons and becomes an anion, with a negative (−) charge.
2. Attraction between the opposite charges then draws the two ions together.
3. An ionic compound is formed.

Figure 2–5a illustrates the formation of an ionic bond. The sodium atom diagrammed in 1 has an atomic number of 11, so this atom normally contains 11 protons and 11 electrons. (Because neutrons are electrically neutral, they do not affect the formation of ions or ionic bonds.) Electrons fill the first and second energy levels, and a single electron occupies the valence shell. Losing that 1 electron would give the sodium atom a full valence shell—the second energy level—and would produce a **sodium ion**, with a charge of +1. (The chemical shorthand for a sodium ion is Na^+.) But a sodium atom cannot simply throw

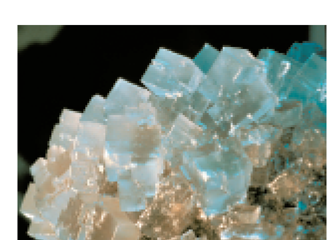
Figure 2–5 The Formation of Ionic Bonds.



a Formation of an ionic bond. 1 A sodium (Na) atom gives up an electron, which is gained by a chlorine (Cl) atom. 2 Because the sodium ion (Na^+) and chloride ion (Cl^-) have opposite charges, they are attracted to one another. 3 The association of sodium and chloride ions forms the ionic compound sodium chloride.



b Sodium chloride crystal. Large numbers of sodium and chloride ions form a crystal of sodium chloride (table salt).



c Photo of sodium chloride crystals

Go to MasteringA&P™ > Study Area > Get Ready for A&P > 2. Get Ready for A&P Website> Chapter 5: Chemistry > Chemical Bonding: Video Tutor: Introduction to Chemical Bonding away the electron: The electron must be donated to an electron acceptor. A chlorine atom has seven electrons in its valence shell, so it needs only one electron to achieve stability. A sodium atom can provide the extra electron. In the process (1), the chlorine atom becomes a **chloride ion** (Cl^-) with a charge of -1.

Both atoms have now become stable ions with filled outermost energy levels. But the two ions do not move apart after the electron transfer, because the positively charged sodium ion is attracted to the negatively charged chloride ion (2). The combination of oppositely charged ions forms an *ionic compound*—in this case, **sodium chloride**, or NaCl (3). Large numbers of sodium and chloride ions interact to form highly structured crystals, held together by the strong electrical attraction of oppositely charged ions (**Figure 2–5b,c**).

Note that ionic compounds are not molecules. That is because they consist of a group of ions rather than atoms bonded by shared electrons. Sodium chloride and other ionic compounds are common in body fluids, but they are not present as intact crystals. When placed in water, many ionic compounds dissolve, and some or all of the component anions and cations separate.

Covalent Bonds

Some atoms can complete their valence shells not by gaining or losing electrons, but by sharing electrons with other atoms. Such sharing creates **covalent** (kō-VĀ-lent) **bonds** between the atoms involved.

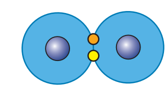
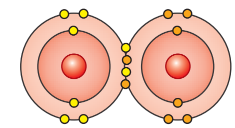
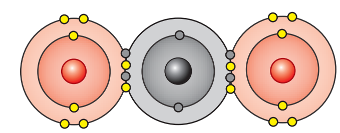
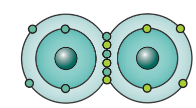
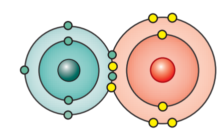
Individual hydrogen atoms, as diagrammed in **Figure 2–2a**, do not exist in nature. Instead, we find hydrogen molecules. Molecular hydrogen consists of a pair of hydrogen atoms (**Figure 2–6**). In chemical shorthand, molecular hydrogen is H_2, where H is the chemical symbol for hydrogen, and the subscript 2 indicates the number of atoms. Molecular hydrogen is a gas that is present in the atmosphere in very small quantities. When the two hydrogen atoms share their electrons, each electron whirls around both nuclei. The sharing of one pair of electrons creates a **single covalent bond**, which can be represented by a single line (—) in the *structural formula* of a molecule.

Oxygen, with an atomic number of 8, has 2 electrons in its first energy level and 6 in its second. The oxygen atoms diagrammed in **Figure 2–6** become stable by sharing two pairs of electrons, forming a **double covalent bond**, represented by two lines (=) in its structural formula. Molecular oxygen (O_2) is an atmospheric gas that most organisms need to survive. Our cells would die without a relatively constant supply of oxygen.

In our bodies, chemical processes that consume oxygen generally also produce **carbon dioxide** (CO_2) as a waste product. Each of the oxygen atoms in a carbon dioxide molecule forms double covalent bonds with the carbon atom, as **Figure 2–6** shows.

A triple covalent bond is the sharing of three pairs of electrons, and is indicated by three lines (≡) in a structural formula. A triple covalent bond joins 2 nitrogen atoms to form molecular nitrogen (N_2) (see **Figure 2–6**). Molecular nitrogen accounts for about 79 percent of our planet's atmosphere, but our cells ignore it completely. In fact, deep-sea divers live for long periods while breathing artificial air that does not contain nitrogen. (We discuss the reasons for eliminating nitrogen under these conditions in the Decompression Sickness Clinical Note in Chapter 23.)

Figure 2–6 Covalent Bonds in Five Common Molecules.

Molecule	Electron-Shell Model and Structural Formula	
Hydrogen (H_2)		H–H
Oxygen (O_2)		O=O
Carbon dioxide (CO_2)		O=C=O
Nitrogen (N_2)		N≡N
Nitric oxide (NO)		N=O

Covalent bonds usually form molecules in which the valence shells of the atoms involved are filled. An atom, ion, or molecule that contains unpaired electrons in its valence shell is called a *free radical*. Free radicals are highly reactive. Almost as fast as it forms, a free radical enters additional reactions that are typically destructive. For example, free radicals can damage or destroy vital compounds, such as proteins. Evidence suggests that the cumulative damage from free radicals inside and outside our cells is a major factor in the aging process. Free radicals sometimes form in the course of normal metabolism, but cells have several methods of removing or inactivating them.

Nitric oxide (NO), however, is a free radical with several important functions in the body (see **Figure 2–6**). It reacts readily with other atoms or molecules, but it is involved in chemical communication in the nervous system, in the control of blood vessel diameter, in blood clotting, and in the defense against bacteria and other pathogens (disease-causing organisms).

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Tips & Tools

Remember this mnemonic for the covalent bonding of hydrogen, oxygen, nitrogen, and carbon atoms: HONC 1234. ***H****ydrogen shares* *1* pair of electrons (H—), ***o****xygen shares* *2* pairs (—O—), ***n****itrogen shares 3 pairs (—N—, with a bond above N), and ***c****arbon shares* *4* pairs (—C—, with bonds above and below C).

Nonpolar Covalent Bonds

Covalent bonds are very strong, because the shared electrons hold the atoms together. In typical covalent bonds the atoms remain electrically neutral, because each shared electron spends just as much time "at home" as away. (If you and a friend were tossing a pair of baseballs back and forth as fast as you could, on average, each of you would have just one baseball.)

Many covalent bonds involve an equal sharing of electrons. Such bonds are called **nonpolar covalent bonds**. They occur, for instance, between 2 atoms of the same type. Nonpolar covalent bonds are very common. In fact, those involving carbon atoms form most of the structural components of the human body.

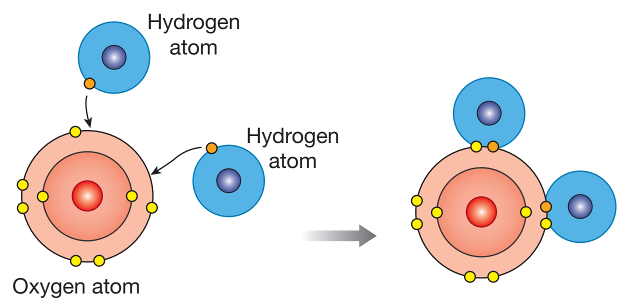
Polar Covalent Bonds

Covalent bonds involving different types of atoms may involve an unequal sharing of electrons, because the elements differ in how strongly they attract electrons. An unequal sharing of electrons creates a **polar covalent bond**. In chemistry, *polarity* refers to a separation of positive and negative electric charge. It can apply to a bond or entire molecule. For example, in a molecule of water, an oxygen atom forms covalent bonds with 2 hydrogen atoms (**Figure 2–7a**). The oxygen nucleus (with its 8 protons) has a much stronger attraction for the shared electrons than the hydrogen atoms (each with a single proton). As a result, the electrons spend more time orbiting the oxygen nucleus than orbiting the hydrogen nuclei.

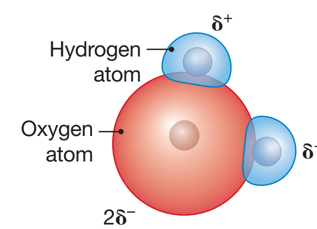
Because the oxygen atom has two extra electrons most of the time, it develops a slight (partial) negative charge, indicated by δ^-, as shown in **Figure 2–7b**. At the same time, each hydrogen atom develops a slight (partial) positive charge, δ^+, because its electron is away much of the time. (Suppose you and a friend were tossing a pair of baseballs back and forth, but one of you returned them as fast as possible while the other held onto them for a while before throwing them back. One of you would now, on average, have more than one baseball, and the other would have less than one.)

The unequal sharing of electrons makes polar covalent bonds somewhat weaker than nonpolar covalent bonds. Polar covalent bonds often create *polar molecules*—molecules that have positive and negative ends. Water is the most important polar molecule in the body.

Figure 2–7 **Water Molecules Contain Polar Covalent Bonds.**



a **Formation of a water molecule**. In forming a water molecule, an oxygen atom completes its valence shell by sharing electrons with a pair of hydrogen atoms. The sharing is unequal, because the oxygen atom holds the electrons more tightly than do the hydrogen atoms.



b **Electric charges on a water molecule**. Because the oxygen atom has 2 extra electrons much of the time, it develops a slight negative charge, and the hydrogen atoms become weakly positive. The bonds in a water molecule are polar covalent bonds.

Hydrogen Bonds

Covalent and ionic bonds tie atoms together to form molecules and/or compounds. Other, comparatively weak forces also act between adjacent molecules, and even between atoms within a large molecule. The most important of these weak attractive forces is the hydrogen bond. A **hydrogen bond** is the attraction between a slight positive charge (δ^+) on the hydrogen atom of a polar covalent bond and a slight negative charge (δ^-) on an oxygen, nitrogen, or fluorine atom of another polar covalent bond.

Hydrogen bonds are too weak to create molecules, but they can change the shapes of molecules or pull molecules closer together. For example, hydrogen bonding occurs between water molecules, forming clumps, or groups, of interconnected water molecules (**Figure 2–8**). At a water surface, this attraction between water molecules slows the rate of evaporation and creates what is known as *surface tension*. Surface tension acts as a barrier that keeps lightweight objects that cannot break through the top layer of water molecules from entering the water. For example, it allows insects to walk across the surface of a pond or puddle. Similarly, the surface tension in a layer of tears on the eye prevents dust particles from touching the surface of the

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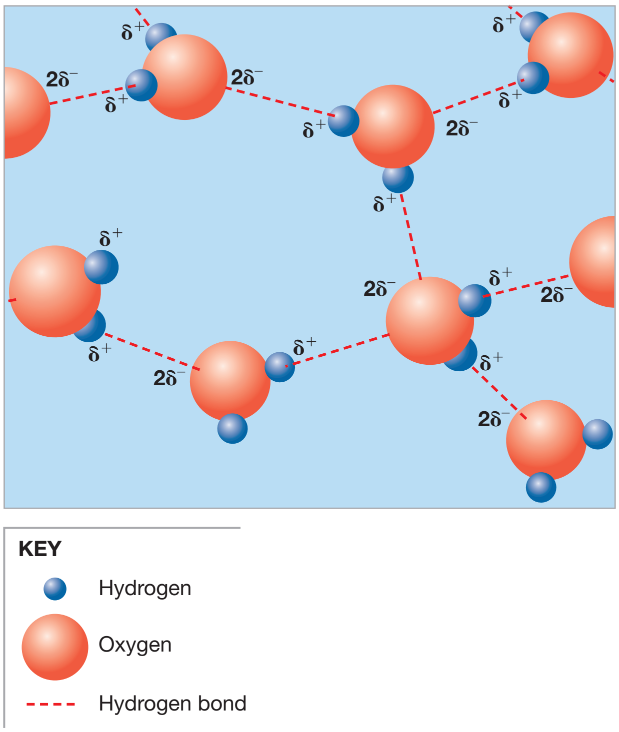
Figure 2–8 Hydrogen Bonds Form between Water Molecules. The hydrogen atoms of a water molecule have a slight positive charge, and the oxygen atom has a slight negative charge (*see Figure 2–7b*). The distances between these molecules have been exaggerated for clarity.



? Hydrogen bonds do not form between two hydrogen atoms. Where do hydrogen bonds form?

eye. At the cellular level, hydrogen bonds affect the shapes and properties of large, complex molecules, such as proteins and nucleic acids (including DNA).

States of Matter

Most matter in our environment exists in one of three states: solid, liquid, or gas. *Solids* keep their volume and their shape at ordinary temperatures and pressures. A lump of granite, a brick, and a textbook are solid objects. *Liquids* have a constant volume, but no fixed shape. The shape of a liquid is determined by the shape of its container. Water, coffee, and soda are liquids. A *gas* has no constant volume and no fixed shape. Gases can be compressed or expanded, and unlike liquids they will fill a container of any size. The most familiar example is the air of our atmosphere.

What determines whether a substance is a solid, liquid, or gas? A matter's state depends on the degree of interaction among its atoms or molecules. The particles of a solid are held tightly together, while those of a gas are very far apart. Water is the only substance that occurs as a solid (ice), a liquid (water), and a gas (water vapor) at temperatures compatible with life. Water exists as a liquid over a broad range of temperatures primarily because of hydrogen bonding among the water molecules. We talk more about water's unusual properties in Section 2-6.

2

√ Checkpoint

4. **Define *chemical bond* and identify several types of chemical bonds.**
5. **Which kind of bond holds atoms in a water molecule together? What attracts water molecules to one another?**
6. **Both oxygen and neon are gases at room temperature. Oxygen combines readily with other elements, but neon does not. Why?**

See the blue Answers tab at the back of the book.

2-3 Decomposition, synthesis, and exchange reactions are important types of chemical reactions in physiology

Learning Outcome Distinguish among the major types of chemical reactions that are important for studying physiology.

Cells stay alive and functional by controlling chemical reactions. In a **chemical reaction**, new chemical bonds form between atoms, or existing bonds between atoms are broken. These changes take place as atoms in the reacting substances, called **reactants**, are rearranged to form different substances, or **products** (see **Spotlight Figure 2–4c**).

In effect, each cell is a chemical factory. Cells use chemical reactions to provide the energy they need to maintain homeostasis and to perform essential functions such as growth, maintenance and repair, secretion (discharging from a cell), and contraction. All of the reactions under way in the cells and tissues of the body at any given moment make up its **metabolism** (me-TAB-ō-lizm).

Basic Energy Concepts

An understanding of some basic relationships between matter and energy is helpful for any discussion of chemical reactions. **Work** is the movement of an object or a change in the physical structure of matter. In your body, work includes movements such as walking or running, and also the synthesis of *organic* (carbon-containing) molecules and the conversion of liquid water to water vapor (evaporation). **Energy** is the capacity to do work, and movement or physical change cannot take place without energy. The two major types of energy are kinetic energy and potential energy:

- **Kinetic energy** is the energy of motion—energy that can be transferred to another object and do work. When you fall off a ladder, it is kinetic energy that does the damage.

2

- **Potential energy** is stored energy—energy that has the potential to do work. It may derive from an object's position (you standing on a ladder) or from its physical or chemical structure (a stretched spring or a charged battery).

Kinetic energy must be used in climbing the ladder, in stretching the spring, or in charging the battery. The resulting potential energy is converted back into kinetic energy when you descend, the spring recoils, or the battery discharges. The kinetic energy can then be used to perform work. For example, in an MP3 player, the chemical potential energy stored in small batteries is converted to kinetic energy that vibrates the sound-producing membranes in headphones or external speakers.

Energy cannot be destroyed: It can only be converted from one form to another. A conversion between potential energy and kinetic energy is never 100 percent efficient. Each time an energy exchange occurs, some of the energy is released in the form of heat. *Heat* is an increase in random molecular motion, and the temperature of an object is proportional to the average kinetic energy of its molecules. Heat can never be completely converted to work or any other form of energy, and cells cannot capture it or use it to do work.

Cells do work as they use energy to synthesize complex molecules and move materials into, out of, and within the cell. The cells of a skeletal muscle at rest, for example, contain potential energy in the form of the positions of protein filaments and the covalent bonds between molecules inside the cells. When a muscle contracts, it performs work. Potential energy is converted into kinetic energy, and heat is released. The amount of heat is proportional to the amount of work done. As a result, when you exercise, your body temperature rises.

Types of Chemical Reactions

Three types of chemical reactions are important to the study of physiology: decomposition reactions, synthesis reactions, and exchange reactions. Many of these chemical reactions are also *reversible reactions*.

Decomposition Reactions

A **decomposition reaction** breaks a molecule into smaller fragments. Here is a diagram of a simple *decomposition reaction*:

$$AB \longrightarrow A + B$$

Decomposition reactions take place outside cells as well as inside them. For example, a typical meal contains molecules of fats, sugars, and proteins that are too large and too complex to be absorbed and used by your body. Decomposition reactions in the digestive tract break these molecules down into smaller fragments that can be absorbed.

Decomposition reactions involving water are important in the breakdown of complex molecules in the body. In this process, which is called a **hydrolysis reaction** (hī-DROL-i-sis; *hydro-*, water + *lysis*, a loosening), one of the bonds in a complex molecule is broken, and the components of a water molecule (H and OH) are added to the resulting fragments:

$$AB + H_2O \longrightarrow AH + BOH$$

Collectively, the decomposition reactions of complex molecules within the body's cells and tissues are referred to as **catabolism** (ka-TAB-ō-lizm; *katabole*, a throwing down). When a covalent bond—a form of potential energy—is broken, it releases kinetic energy that can do work. By harnessing the energy released in this way, cells carry out vital functions such as growth, movement, and reproduction.

Synthesis Reactions

Synthesis (SIN-the-sis) is the opposite of decomposition. A **synthesis reaction** assembles smaller molecules into larger molecules. A simple synthetic reaction is diagrammed here:

$$A + B \longrightarrow AB$$

Synthesis reactions may involve individual atoms or the combination of molecules to form even larger products. The formation of water from hydrogen and oxygen molecules is a synthesis reaction. Synthesis always involves the formation of new chemical bonds, whether the reactants are atoms or molecules.

A **dehydration synthesis**, or *condensation*, **reaction** is the formation of a complex molecule by the removal of a water molecule:

$$AH + BOH \longrightarrow AB + H_2O$$

Dehydration synthesis is the opposite of hydrolysis. We look at examples of both reactions in later sections (Sections 2-9, 2-10, and 2-11).

Collectively, the synthesis of new molecules within the body's cells and tissues is known as **anabolism** (a-NAB-ō-lizm; *anabole*, a throwing upward). Anabolism is usually considered an "uphill" process because it takes energy to create a chemical bond (just as it takes energy to push something uphill). Cells must balance their energy budgets, with catabolism providing the energy to support anabolism and other vital functions.

Tips & Tools

To remember the difference between anabolism (synthesis) and catabolism (breakdown), relate the terms to words you already know: *Anabolic* steroids are used to build up muscle tissue, while both *catastrophe* and *catabolism* involve destruction (breakdown).

Exchange Reactions

In an **exchange reaction**, parts of the reacting molecules are shuffled around to produce new products:

$$AB + CD \longrightarrow AD + CB$$

The reactants and products contain the same components (A, B, C, and D), but those components are present in different combinations. In an exchange reaction, the reactant molecules AB and CD must break apart (a decomposition) before they can interact with each other to form AD and CB (a synthesis).

Reversible Reactions

At least in theory, chemical reactions are reversible, so if $A + B \longrightarrow AB$, then $AB \longrightarrow A + B$. Many important biological reactions are freely reversible. Such reactions can be represented as an equation:

$$A + B \rightleftharpoons AB$$

This equation indicates that, in a sense, two reactions are taking place at the same time. One is a synthesis reaction $(A + B \rightarrow AB)$ and the other is a decomposition reaction $(AB \rightarrow A + B)$.

Recall from Chapter 1 that a state of *equilibrium* exists when opposing processes or forces are in balance. At equilibrium, the rates of the two reactions are in balance. As fast as one molecule of AB forms, another degrades into A + B.

What happens when equilibrium is disturbed—say, if you add more AB? In our example, the rate of the synthesis reaction is directly proportional to the frequency of encounters between A and B. In turn, the frequency of encounters depends on the degree of crowding. (You are much more likely to bump into another person in a crowded room than in a room that is almost empty.) So adding more AB molecules will increase the rate of conversion of AB to A and B. The amounts of A and B will then increase, leading to an increase in the rate of the reverse reaction—the formation of AB from A and B. Eventually, a balance, or equilibrium, is again established.

Tips & Tools

Jell-O provides an example of a physical reversible reaction. Once Jell-O has been refrigerated, the gelatin sets up and forms a solid, but if it sits without refrigeration for too long, it turns back into a liquid again.

Checkpoint

7. **Using the rules for chemical notation, how is an ion's electrical charge represented?**
8. **Using the rules for chemical notation, write the molecular formula for glucose, a compound composed of 6 carbon atoms, 12 hydrogen atoms, and 6 oxygen atoms.**
9. **Identify and describe three types of chemical reactions important to human physiology.**
10. **In cells, glucose, a six-carbon molecule, is converted into two three-carbon molecules by a reaction that releases energy. What type of reaction is this?**

See the blue Answers tab at the back of the book.

2-4 Enzymes speed up reactions by lowering the energy needed to start them

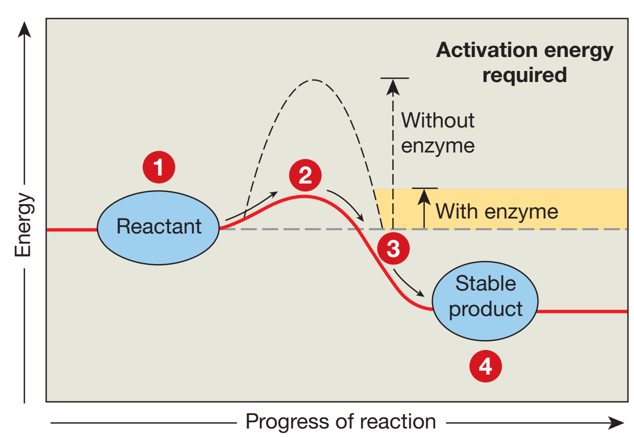
Learning Outcome Describe the crucial role of enzymes in metabolism.

Most **biochemical reactions** (those that happen in living organisms) do not take place spontaneously, or if they do, they occur so slowly that they would be of little value to living cells. Before a reaction can proceed, enough energy must be provided to activate the reactants. The amount of energy required to start a reaction is called the **activation energy**. Many reactions can be activated by changes in temperature or acidity, but such changes are deadly to cells. For example, every day your cells break down complex sugars as part of your normal metabolism. Yet to break down a complex sugar in a laboratory, you must boil it in an acidic solution. Your cells don't have that option! Temperatures that high and solutions that corrosive would immediately destroy living tissues. Instead, your cells use special proteins called *enzymes* to catalyze (speed up) most of the complex synthesis and decomposition reactions in your body.

Enzymes promote chemical reactions by lowering their required activation energy (Figure 2–9). In doing so, they make it possible for chemical reactions, such as the breakdown of sugars, to proceed under conditions compatible with life. Cells make enzyme molecules, each of which promotes a specific reaction.

Enzymes belong to a class of substances called **catalysts** (KAT-uh-lists; *katalysis*, dissolution), compounds that speed up

Figure 2–9 **Enzymes Lower Activation Energy.** Enzymes lower the activation energy required for a chemical reaction to proceed readily (in order, from 1–4) under conditions in the body.



? Which number represents the greatest amount of energy that must be overcome during the reaction? Which number represents the lowest amount of reaction energy?

2

chemical reactions without themselves being permanently changed or consumed. Enzymatic reactions, which are reversible, can be written as

$$A + B \overset{\text{enzyme}}{\rightleftharpoons} AB$$

An appropriate enzyme can accelerate, or speed up, a reaction, but an enzyme affects only the *rate* of the reaction, not its direction or the products that are formed. An enzyme cannot bring about a reaction that would otherwise be impossible. Enzymatic reactions are generally reversible, and they proceed until equilibrium is reached.

The complex reactions that support life take place in a series of interlocking steps, each controlled by a specific enzyme. Such a reaction sequence is called a *metabolic pathway*. A synthetic pathway can be diagrammed as

$$A \xrightarrow[\text{Step 1}]{\text{enzyme 1}} B \xrightarrow[\text{Step 2}]{\text{enzyme 2}} C \xrightarrow[\text{Step 3}]{\text{enzyme 3}} \text{and so on.}$$

In many cases, the steps in the synthetic pathway differ from those in the decomposition pathway, and separate enzymes are often involved.

It takes activation energy to start a chemical reaction, but once it has begun, the reaction as a whole may absorb or release energy as it proceeds to completion. If the amount of energy released is greater than the activation energy needed to start the reaction, there will be a net release of energy. Reactions that release energy are said to be **exergonic** (*exo-*, outside + *ergon*, work). Exergonic reactions are relatively common in the body. They generate the heat that maintains your body temperature.

If more energy is required to begin the reaction than is released as it proceeds, the reaction as a whole will absorb energy. Such reactions are called **endergonic** (*endo-*, inside). The synthesis of molecules such as fats and proteins results from endergonic reactions.

√ Checkpoint

11. **What is an enzyme?**
12. **Why are enzymes needed in our cells?**

See the blue Answers tab at the back of the book.

2-5 Inorganic compounds lack carbon, and organic compounds contain carbon

Learning Outcome Distinguish between inorganic compounds and organic compounds.

The human body is very complex, but it contains relatively few elements (see **Table 2–1**, p. 29). Just knowing the identity and quantity of each element in the body will not help you understand the body any more than memorizing the alphabet will help you understand this text. Just as 26 letters can be combined to form thousands of different words in this text, only about 26 elements combine to form thousands of different chemical compounds in our bodies. As we saw in Chapter 1, these compounds make up the living cells that form the body's framework and carry on all its life processes. Learning about the major classes of chemical compounds will help you to understand the structure and function of the human body.

Two of the major classes of compounds are nutrients and metabolites. **Nutrients** are the substances from food that are necessary for normal physiological functions. Nutrients include carbohydrates, proteins, lipids (fats), vitamins, minerals, and water. **Metabolites** (me-TAB-ō-lītz; *metabole*, change) are substances that are involved in, or are a by-product of, metabolism. We can broadly categorize nutrients and metabolites as either inorganic or organic. **Inorganic compounds** generally do not contain carbon and hydrogen atoms as their primary structural components. (If present, they do not form C—H bonds.) In contrast, carbon and hydrogen always form the basis for **organic compounds**. Their molecules can be much larger and more complex than inorganic compounds. *Carbohydrates*, *proteins*, and *lipids* are organic nutrients used by the body—we cover these in later sections.

The most important inorganic compounds in the body are as follows:

- *carbon dioxide*, a by-product of cell metabolism;
- *oxygen*, an atmospheric gas required in important metabolic reactions;
- *water*, which accounts for more than half of our body weight; and
- *acids*, *bases*, and *salts*—compounds held together partially or completely by ionic bonds.

In the next section, we focus on water, its properties, and how those properties establish the conditions necessary for life. Most of the other inorganic molecules and compounds in the body exist in association with water, the primary component of our body fluids. Both carbon dioxide and oxygen, for example, are gas molecules that are transported in body fluids. Also, all the inorganic acids, bases, and salts we will discuss are dissolved in body fluids.

√ Checkpoint

13. **Compare inorganic compounds to organic compounds.**

See the blue Answers tab at the back of the book.

2-6 Physiological systems depend on water

Learning Outcome Explain how the chemical properties of water make life possible.

Water (H_2O) is the most important substance in the body. It makes up to two-thirds of total body weight. A change in the body's water content can be fatal, because virtually all physiological systems will be affected.

Although water is familiar to everyone, it has some highly unusual properties: universal solvent, reactivity, high heat capacity, and lubrication. All are important to the human body.

- *Universal Solvent*. A remarkable number of inorganic and organic molecules and compounds are water *soluble*, meaning they will dissolve or break up in water. The individual particles become distributed within the water, and the result is a **solution**—a uniform mixture of two or more substances. The liquid in which other atoms, ions, or molecules are distributed is called the **solvent**. The dissolved substances are the **solutes**. In *aqueous* (AK-wē-us) *solutions*, water is the solvent. Water is often called a "universal solvent" because more substances dissolve in it than any other liquid.
- *Reactivity*. In our bodies, chemical reactions take place in water, but water molecules are also reactants in some reactions. Hydrolysis and dehydration synthesis are two examples noted earlier in the chapter.
- *High Heat Capacity*. **Heat capacity** is the quantity of heat required to raise the temperature of a unit mass of a substance 1°C. Water has an unusually high heat capacity, because water molecules in the liquid state are attracted to one another through hydrogen bonding. Important consequences of this attraction include the following:
  - The temperature of water must be quite high (it requires a lot of energy) to break all of the hydrogen bonds between individual water molecules, and allow them to escape and become water vapor, a gas. Therefore, water remains a liquid over a broad range of environmental temperatures, and the freezing and boiling points of water are far apart.
  - Water carries a great deal of heat away with it when it changes from a liquid to a gas. This feature explains the cooling effect of perspiration on the skin.
  - An unusually large amount of heat energy is required to change the temperature of 1 g of water by 1°C. As a result, a large mass of water changes temperature slowly. This property is called *thermal inertia*. Thermal inertia helps stabilize body temperature because water accounts for up to two-thirds of the weight of the human body.
- *Lubrication*. Water is an effective lubricant because there is little friction between water molecules. So even a thin layer of water between two opposing surfaces will greatly reduce friction between them. (That is why driving on wet roads can be tricky. Your tires may start sliding on a layer of water rather than maintaining contact with the road.) Within joints such as the knee, an aqueous solution prevents friction between the opposing surfaces. Similarly, a small amount of fluid in the body cavities prevents friction between internal organs, such as the heart or lungs, and the body wall. ⮌ p. 15

The Properties of Aqueous Solutions

2

Water's chemical structure makes it an unusually effective solvent. The covalent bonds in a water molecule are oriented so that the hydrogen atoms are fairly close together. As a result, the water molecule has positive- and negative-charged ends, or poles (**Figure 2–10a**). For this reason, a water molecule is called a **polar molecule**.

Many inorganic compounds are held together partly or completely by ionic bonds. In water, these compounds undergo **dissociation** (di-sō-sē-Ā-shun)—the splitting of a compound into smaller molecules. **Ionization** (ī-on-i-ZĀ-shun) is the dissociation into ions. In this process, ionic bonds are broken as the individual ions interact with the positive or negative ends of polar water molecules (**Figure 2–10b**). The result is a mixture of cations and anions surrounded by water molecules. The water molecules around each ion form a *hydration sphere* that isolates the ions from each other, thus preventing the formation of ionic bonds.

An aqueous solution containing anions and cations will conduct an electrical current. When this happens, cations $(+)$ move toward the negative side, and anions $(-)$ move toward the positive side. Electrical forces across plasma membranes affect the functioning of all cells, and small electrical currents carried by ions are essential to muscle contraction and nerve function. (We will discuss these processes in more detail in Chapters 10 and 12.)

Electrolytes and Body Fluids

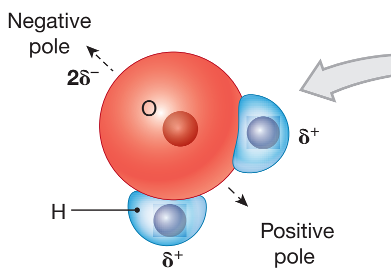
Soluble inorganic substances whose ions will conduct an electrical current in solution are called **electrolytes** (e-LEK-trō-lītz). Sodium chloride in solution is an electrolyte. The dissociation of electrolytes in blood and other body fluids releases a variety of ions. **Table 2–2** lists important electrolytes and the ions released when they dissociate.

Changes in the concentrations of electrolytes in body fluids will disturb almost every vital function. For example, a declining potassium ion (K^+) level will lead to a general muscular paralysis, and a rising concentration will cause weak and irregular heartbeats. The concentrations of ions in body fluids are

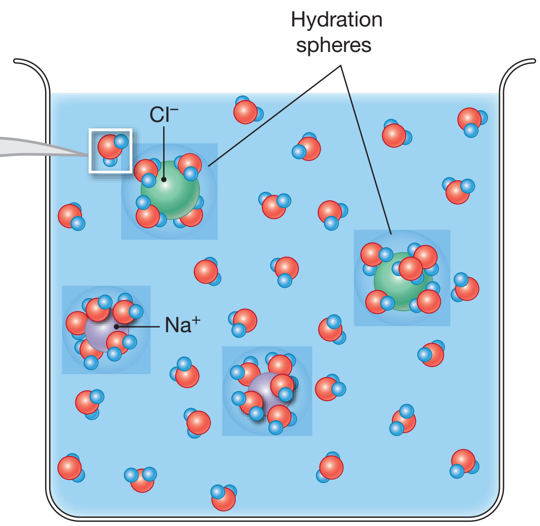
Table 2–2 Important Electrolytes That Dissociate in Body Fluids

Electrolyte	Ions Released
$NaCl$ (sodium chloride)	$\rightarrow Na^+ + Cl^-$
KCl (potassium chloride)	$\rightarrow K^+ + Cl^-$
$CaPO_4$ (calcium phosphate)	$\rightarrow Ca^{2+} + PO_4{}^{2-}$
$NaHCO_3$ (sodium bicarbonate)	$\rightarrow Na^+ + HCO_3{}^-$
$MgCl_2$ (magnesium chloride)	$\rightarrow Mg^{2+} + 2Cl^-$
Na_2HPO_4 (sodium hydrogen phosphate)	$\rightarrow 2Na^+ + HPO_4{}^{2-}$
Na_2SO_4 (sodium sulfate)	$\rightarrow 2Na^+ + SO_4{}^{2-}$

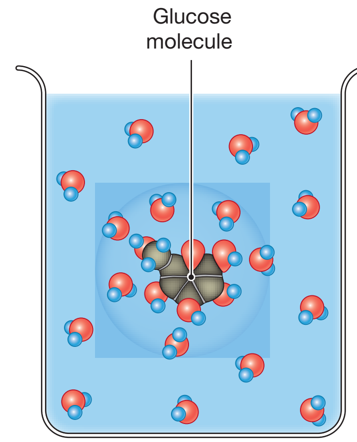
Figure 2–10 **Water Molecules Surround Solutes in Aqueous Solutions.**



a **Water molecule**. In a water molecule, oxygen forms polar covalent bonds with 2 hydrogen atoms. Because both hydrogen atoms are at one end of the molecule, it has an uneven distribution of electric charges, creating positive and negative poles.



b **Sodium chloride in solution**. Ionic compounds, such as sodium chloride, dissociate in water as the polar water molecules break the ionic bonds in the large crystal structure. Each ion in solution is surrounded by water molecules, creating hydration spheres.



c **Glucose in solution**. Hydration spheres also form around an organic molecule containing polar covalent bonds. If the molecule binds water strongly, as does glucose, it will be carried into solution—in other words, it will dissolve. Note that the molecule does not dissociate, as occurs for ionic compounds.

carefully regulated, mostly by the coordination of activities at the kidneys (ion excretion), the digestive tract (ion absorption), and the skeletal system (ion storage or release).

Hydrophilic and Hydrophobic Compounds

Some organic molecules contain polar covalent bonds, which also attract water molecules. The hydration spheres that form may then carry these molecules into solution. Molecules that interact readily with water molecules in this way are called **hydrophilic** (hī-drō-FIL-ik; *hydro-*, water + *philos*, loving). Glucose, an important soluble sugar, is one example (Figure 2–10c).

When nonpolar molecules are exposed to water, hydration spheres do not form and the molecules do not dissolve. Molecules that do not readily interact with water are called **hydrophobic** (hī-drō-FŌB-ik; *hydro-*, water + *phobos*, fear). Fats and oils of all kinds are some of the most familiar hydrophobic molecules. For example, body fat deposits consist of large, hydrophobic droplets trapped in the watery interior of cells. Gasoline and heating oil are hydrophobic molecules not found in the body. When accidentally spilled into lakes or oceans, they form long-lasting oil slicks instead of dissolving.

Tips & Tools

To distinguish between hydrophobic and hydrophilic, remember that a phobia is a fear of something, and that *-philic* ends with "lic," which resembles "like."

Colloids and Suspensions

Body fluids may contain large and complex organic molecules, such as proteins, that are held in solution by their association with water molecules (see Figure 2–10c). A solution containing dispersed proteins or other large molecules is called a **colloid** (KOL-oyd). Liquid Jell-O is a familiar colloid. The particles or molecules in a colloid will remain in solution indefinitely.

In contrast, a **suspension** contains large particles in solution, but if undisturbed, its particles will settle out of solution due to the force of gravity. For example, stirring beach sand into a bucket of water creates a temporary suspension that will last only until the sand settles to the bottom. Whole blood is another temporary suspension, because the blood cells are suspended in the blood plasma. If clotting is prevented, the cells in a blood sample will gradually settle to the bottom of the container. Measuring that settling rate, or "sedimentation rate," is a common laboratory test.

✓ Checkpoint

14. **Explain how the chemical properties of water make life possible.**

See the blue Answers tab at the back of the book.

2-7 Body fluid pH is vital for homeostasis

Learning Outcome Explain what pH is and discuss its importance.

A hydrogen atom involved in a chemical bond or participating in a chemical reaction can easily lose its electron to become a **hydrogen ion (H^+)**. Hydrogen ions are extremely reactive in solution. In excessive numbers, they will disrupt cell and tissue functions. As a result, the concentration of hydrogen ions in body fluids must be regulated precisely.

A few hydrogen ions are normally present even in a sample of pure water, because some of the water molecules dissociate spontaneously, releasing cations and anions. The dissociation of water is a reversible reaction. We can represent it as

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

Notice that the dissociation of one water molecule yields a hydrogen ion (H^+) and a **hydroxide** (hī-DROK-sīd) **ion (OH^-)**.

However, very few water molecules ionize in pure water, so the number of hydrogen and hydroxide ions is small. The quantities are usually reported in moles, making it easy to keep track of the numbers of hydrogen and hydroxide ions. One liter of pure water contains about 0.0000001 mol of hydrogen ions and an equal number of hydroxide ions. In other words, the concentration of hydrogen ions in a solution of pure water is 0.0000001 mol per liter. This can be written as

$$[H^+] = 1 \times 10^{-7} \text{mol/L}$$

The brackets around the H^+ signify "the concentration of," another example of chemical notation.

The hydrogen ion concentration in body fluids is so important to physiological processes that we use a special shorthand to express it. The **pH** of a solution is defined as the negative logarithm of the hydrogen ion concentration in moles per liter. So instead of using the equation $[H^+] = 1 \times 10^{-7}$ mol/L, we say that the pH of pure water is $-(-7)$, or 7.

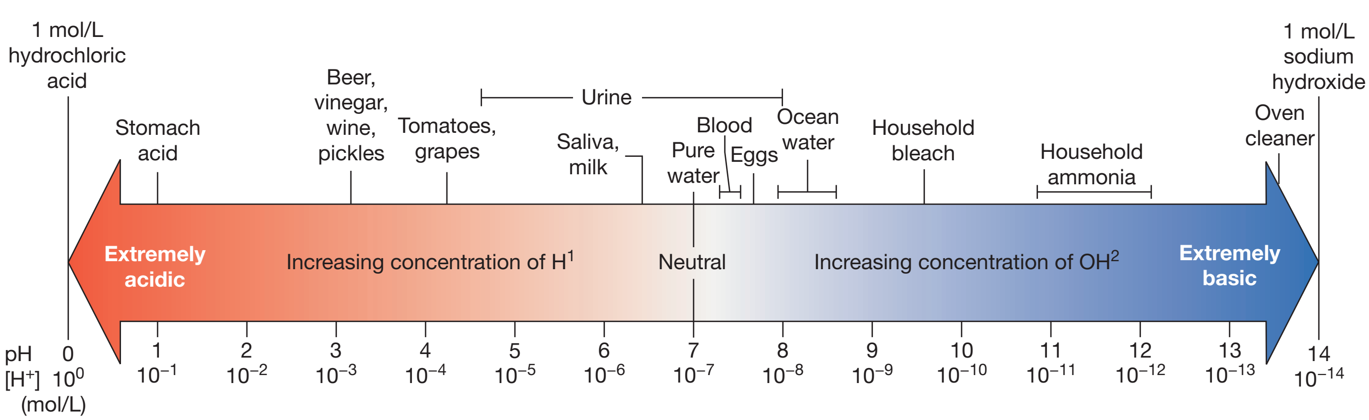
Using pH values saves space, but always remember that the pH number is an *exponent* and that the pH scale is logarithmic. For instance, a pH of 6 ($[H^+] = 1 \times 10^{-6}$, or 0.000001 mol/L) means that the concentration of hydrogen ions is *10 times greater than* it is at a pH of 7 ($[H^+] = 1 \times 10^{-7}$, or 0.0000001 mol/L). The pH scale ranges from 0 to 14 (**Figure 2–11**).

Pure water has a pH of 7, but as **Figure 2–11** indicates, solutions display a wide range of pH values, depending on the nature of the solutes involved:

- A solution with a pH of 7 is said to be **neutral**, because it contains equal numbers of hydrogen and hydroxide ions.
- A solution with a pH below 7 is **acidic** (a-SI-dik), meaning that it contains more hydrogen ions than hydroxide ions.
- A solution with a pH above 7 is **basic**, or *alkaline* (AL-kuh-lin), meaning that it has more hydroxide ions than hydrogen ions.

The normal pH of blood ranges from 7.35 to 7.45. Abnormal fluctuations in pH can damage cells and tissues by breaking chemical bonds, changing the shapes of proteins, and altering cellular functions. *Acidosis* is an abnormal physiological state caused by low blood pH (below 7.35). A pH below 7 can produce coma. Likewise, *alkalosis* results from an abnormally high pH (above 7.45). A blood pH above 7.8 generally causes uncontrollable and sustained skeletal muscle contractions.

Figure 2–11 The pH Scale Indicates Hydrogen Ion Concentration. The pH scale is logarithmic; an increase or decrease of one unit corresponds to a tenfold change in H^+ concentration.



2

Checkpoint

15. **Define pH, and explain how the pH scale relates to acidity and alkalinity.**
16. **What is the significance of pH in physiological systems?**

See the blue Answers tab at the back of the book.

2-8 Acids, bases, and salts have important physiological roles

Learning Outcome Describe the physiological roles of acids, bases, and salts and the role of buffers in body fluids.

Acids and Bases

The body contains *acids* and *bases* that may cause acidosis or alkalosis, respectively. An **acid** (*acere*, sour) is any solute that dissociates in solution and releases hydrogen ions, lowering the pH. A hydrogen atom that loses its electron consists solely of a proton, so we often refer to hydrogen ions simply as protons, and to acids as *proton donors*. In contrast, a **base** is a solute that removes hydrogen ions from a solution, raising the pH. It acts as a *proton acceptor*. Acids and bases are often identified in terms of *strength*. Strength is a measure of the degree of dissociation of either an acid or a base when it is in solution. There are both strong and weak acids and bases.

A *strong acid* dissociates completely in solution, and the reaction occurs essentially in one direction only. *Hydrochloric acid* (HCl) is a representative strong acid. In water, it ionizes as follows:

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

The stomach produces this powerful acid to help break down food. Hardware stores sell HCl under the name muriatic acid, for cleaning concrete and swimming pools.

In solution, many bases release a hydroxide ion (OH^-). Hydroxide ions have an attraction for hydrogen ions and react quickly with them to form water molecules. A *strong base* dissociates completely in solution. *Sodium hydroxide*, NaOH, is a strong base. In solution, it releases sodium ions and hydroxide ions:

$$NaOH \longrightarrow Na^{+} + OH^{-}$$

Strong bases have a variety of industrial and household uses. Drain openers (such as Drāno) and lye are two familiar examples.

Weak acids and *weak bases* do not dissociate completely. At equilibrium, a significant number of molecules remain intact in the solution. For the same number of molecules in solution, weak acids and weak bases have less impact on pH than do strong acids and strong bases. *Carbonic acid* (H_2CO_3) is a weak acid found in body fluids. In solution, carbonic acid reversibly dissociates into a hydrogen ion and a *bicarbonate ion*, HCO_3^-:

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

Salts

A **salt** is an ionic compound containing any cation except a hydrogen ion, and any anion except a hydroxide ion. Because they are held together by ionic bonds, many salts dissociate completely in water, releasing cations and anions. For example, sodium chloride (NaCl) dissociates immediately in water, releasing Na^+ and Cl^-. Sodium and chloride ions are the most abundant ions in body fluids. However, many other ions are present in lesser amounts as a result of the dissociation of other ionic compounds. Ionic concentrations in the body are regulated in ways we describe in Chapters 26 and 27.

The dissociation of sodium chloride does not affect the local concentrations of hydrogen ions or hydroxide ions. For this reason, NaCl, like many salts, is a "neutral" solute. It does not make a solution more acidic or more basic. In contrast, some salts may interact with water molecules and indirectly affect the concentrations of H^+ and OH^- ions. Thus, in some cases, the dissociation of salts makes a solution slightly acidic or slightly basic.

Buffers and pH Control

Buffers are compounds that stabilize the pH of a solution by removing or replacing hydrogen ions. *Buffer systems* usually involve a weak acid and its related salt, which functions as a weak base. For example, the carbonic acid–bicarbonate buffer system (detailed in Chapter 27) consists of carbonic acid and sodium bicarbonate, $NaHCO_3$, otherwise known as baking soda. Buffers and buffer systems in body fluids help maintain the pH within normal limits. The pH of several body fluids is included in **Figure 2–11**.

The use of antacids is one example of the type of reaction that takes place in buffer systems. Antacids use sodium bicarbonate to neutralize excess hydrochloric acid in the stomach.

Note that the effects of neutralization are most evident when you add a strong acid to a strong base. For example, by adding hydrochloric acid to sodium hydroxide, you neutralize both the strong acid and the strong base:

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

This neutralization reaction produces water and a salt—in this case, the neutral salt sodium chloride.

Checkpoint

17. **Define the following terms: *acid*, *base*, and salt.**
18. **How does an antacid help decrease stomach discomfort?**

See the blue Answers tab at the back of the book.

2-9 Living things contain organic compounds made up of monomers, polymers, and functional groups

Learning Outcome: Describe monomers and polymers, and the importance of functional groups in organic compounds.

The macromolecules of living things are all organic compounds. Many of these large organic molecules are made up of long chains of carbon atoms linked by covalent bonds. The carbon atoms typically form additional covalent bonds with hydrogen or oxygen atoms and, less commonly, with nitrogen, phosphorus, sulfur, iron, or other elements.

The *macromolecules* of life are complex structures with varied functions and properties. They include carbohydrates, lipids, proteins, and nucleic acids. Each macromolecule is made up of monomer subunits. A **monomer** (MON-ō-mer; *mono-* single + *-mer*, member of a group) is a molecule that can be bonded to other identical molecules to form a **polymer**. Repeating monomers join together through *dehydration synthesis* reactions to form polymers, sometimes called *mers*.

The monomers of carbohydrates, lipids, proteins, and nucleic acids are separated, or released, through *hydrolysis reactions*. As a result of such reactions, carbohydrates release monosaccharides, lipids (fats) release fatty acids and glycerol, proteins release amino acids, and nucleic acids release nucleotides. We discuss the structures and functions of these molecules in the next four sections.

Organic compounds are diverse, but certain groupings of atoms, known as *functional groups*, are responsible for the characteristic reactions of a particular compound. Furthermore, these functional groups greatly influence the properties of any molecule of which they are a part. Table 2–3 details the functional groups you will study in this chapter.

Checkpoint

19. What macromolecules are important to living things?
20. Which functional group acts as an acid?

See the blue Answers tab at the back of the book.

2-10 Carbohydrates contain carbon, hydrogen, and oxygen in a 1:2:1 ratio

Learning Outcome Discuss the structures and functions of carbohydrates.

A **carbohydrate** is an organic molecule that contains carbon, hydrogen, and oxygen in a ratio near 1:2:1. Familiar carbohydrates include the sugars and starches that make up about half of the typical U.S. diet. Carbohydrates typically account for less than 1 percent of total body weight. Carbohydrates are most important as energy sources that are catabolized. In the following sections, we focus on *monosaccharides*, *disaccharides*, and *polysaccharides*.

Monosaccharides

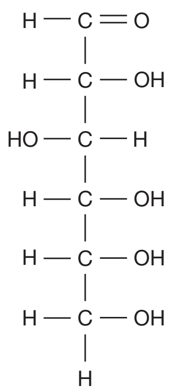
A **monosaccharide** (mon-ō-SAK-uh-rīd; *mono-*, single + *sakcharon*, sugar), or *simple sugar*, is a carbohydrate with three to seven carbon atoms. Depending on how many carbons it contains, a monosaccharide can be called a *triose* (three carbon atoms), *tetrose* (four carbon atoms), *pentose* (five carbon atoms), *hexose* (six carbon atoms), or *heptose* (seven carbon atoms). The *-ose* ending indicates a sugar. The hexose **glucose** (GLŪ-kōs), $C_6H_{12}O_6$, is the most important metabolic "fuel" in the body. Monosaccharides such as glucose dissolve readily in water and are rapidly distributed throughout the body by blood and other body fluids.

Table 2–3 Important Functional Groups of Organic Compounds

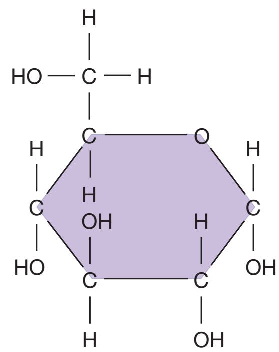
Functional Group	Structural Formula*	Importance	Examples
Amino group —NH_2	R—N(—H)—H	Acts as a base, accepting H^+, depending on pH; can form bonds with other molecules	Amino acids
Carboxyl group —COOH	R—C(—OH)═O	Acts as an acid, releasing H^+ to become R—COO^-	Fatty acids, amino acids
Hydroxyl group —OH	R—O—H	May link molecules through dehydration synthesis (condensation); hydrogen bonding between hydroxyl groups and water molecules; affects solubility	Carbohydrates, fatty acids, amino acids
Phosphate group —PO_4	R—O—P(═O)(—O^-)—O^-	May link other molecules to form larger structures; may store energy in high-energy bonds	Phospholipids, nucleic acids, high-energy compounds

*A structural formula shows the covalent bonds within a molecule or functional group (see Figure 2–6). The letter R represents the term *R group* and is used to denote the rest of the molecule to which a functional group is attached.

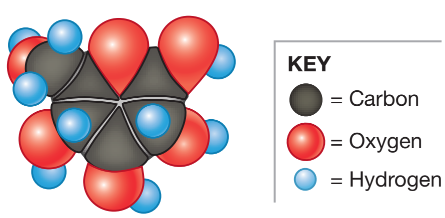
Figure 2–12 The Structures of Glucose. Note that the ring form, the most common form of glucose, is represented with a kind of shorthand in later figures: We leave out the carbon atom at five corners of the hexagon, although we do show the oxygen atom at the remaining corner.



a The structural formula of the straight-chain form



b The structural formula of the ring form, the most common form of glucose



c A three-dimensional model showing the organization of atoms in the ring form

? How many oxygen atoms are shown in each glucose structure?

As shown by two-dimensional diagrams known as *structural formulas*, the atoms in a glucose molecule may form either a straight chain (**Figure 2–12a**) or a ring (**Figure 2–12b**). In the body, the ring form is more common. A three-dimensional model shows the arrangement of atoms in the ring form most accurately (**Figure 2–12c**).

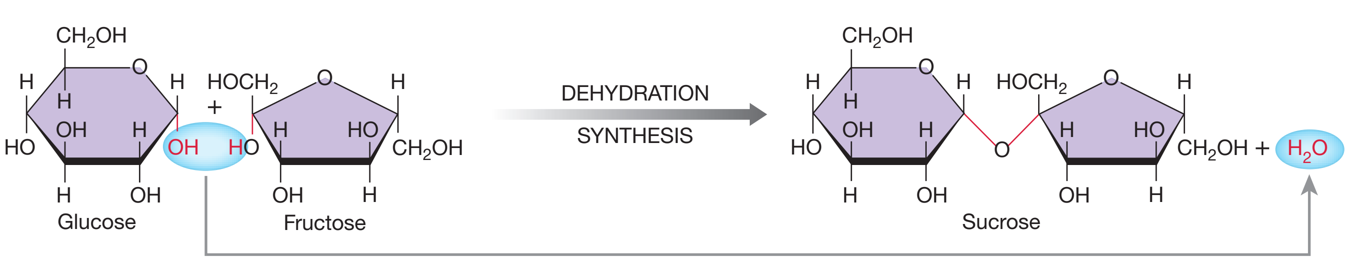
The three-dimensional structure of an organic molecule is an important characteristic, because it usually determines the molecule's fate or function. Some molecules have the same molecular formula—in other words, the same types and numbers of atoms—but different structures. Such molecules are called **isomers**. The body usually treats different isomers as distinct molecules. For example, the monosaccharides glucose and fructose are isomers. *Fructose* is a hexose found in many fruits and in secretions of the male reproductive tract. It has the same chemical formula as glucose, $C_6H_{12}O_6$, but the arrangement of its atoms differs from that of glucose. As a result, separate enzymes and reaction sequences control its breakdown and synthesis.

Disaccharides and Polysaccharides

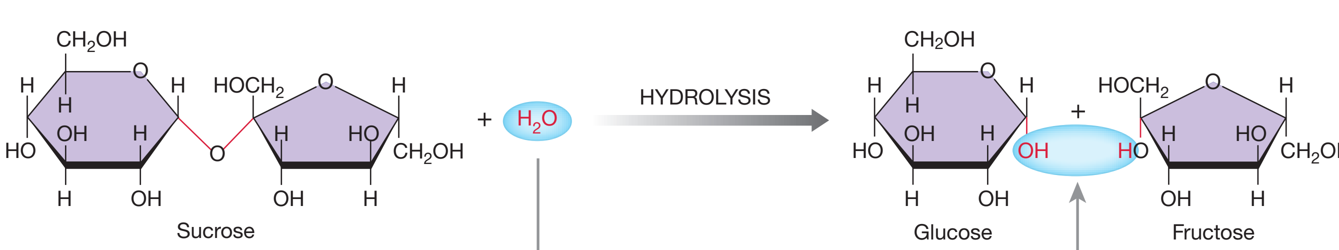
Carbohydrates other than simple sugars are complex molecules composed of monosaccharide building blocks, or monomers. Two monosaccharide monomers joined together form a **disaccharide** (dī-SAK-uh-rīd; *di-*, two). Disaccharides such as *sucrose* (table sugar) have a sweet taste and, like monosaccharides, are quite soluble in water.

The formation of sucrose involves a dehydration synthesis reaction (**Figure 2–13a**). Recall that dehydration synthesis reactions link molecules together by the removal of a water

Figure 2–13 The Formation and Breakdown of Complex Sugars.



a Formation of the disaccharide sucrose through dehydration synthesis. During dehydration synthesis, 2 molecules are joined by the removal of a water molecule.



b Breakdown of sucrose into simple sugars by hydrolysis. Hydrolysis reverses the steps of dehydration synthesis; a complex molecule is broken down by the addition of a water molecule.

molecule. The breakdown of sucrose into simple sugars is an example of hydrolysis, or breakdown by the addition of a water molecule (Figure 2–13b). Hydrolysis is the functional opposite of dehydration synthesis.

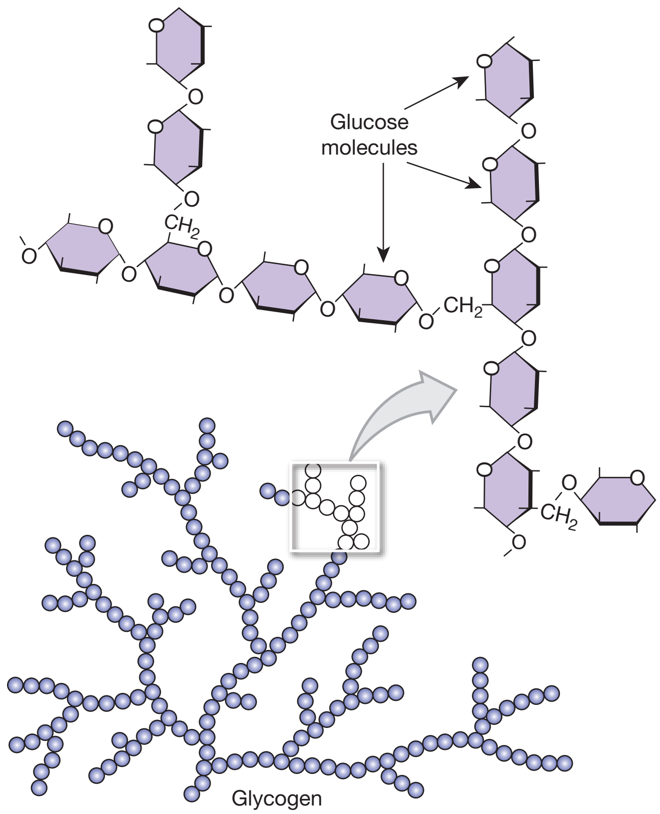
Many foods contain disaccharides, but all carbohydrates except monosaccharides must be broken apart through hydrolysis before they can provide useful energy. Most popular junk foods (high in calories but otherwise lacking in nutritional content), such as candies and sodas, are full of monosaccharides (commonly fructose) and disaccharides (generally sucrose). Some people cannot tolerate sugar for medical reasons. Others avoid it in an effort to control their weight (because excess sugars are converted to fat for long-term storage). Many of these people use *artificial sweeteners* in their foods and beverages. These compounds have a very sweet taste, but they either cannot be broken down in the body or are used in insignificant amounts.

More complex carbohydrates result when repeated dehydration synthesis reactions add additional monosaccharides or disaccharides. These large molecules (polymers) are called **polysaccharides** (pol-ē-SAK-uh-rīdz; *poly-*, many). Polysaccharide chains can be straight or highly branched. *Cellulose*, a structural component of many plants, is a polysaccharide that our bodies cannot digest because the particular linkages between the glucose molecules cannot be cleaved by enzymes in the body. Foods such as celery, which contains cellulose, water, and little else, contribute fiber to digestive wastes but do not provide a source of energy.

Starches are large polysaccharides formed from glucose molecules. Most starches are manufactured by plants. Your digestive tract can break these molecules into monosaccharides. Starches such as those in potatoes and grains are a major dietary energy source.

The polysaccharide **glycogen** (GLĪ-kō-jen), or *animal starch*, has many side branches consisting of chains of glucose molecules (Figure 2–14). Like most other starches, glycogen does not dissolve in water or other body fluids. Muscle cells make and store glycogen. When muscle cells have a high demand for glucose, glycogen molecules are broken down. When the need is low, these cells absorb glucose from the bloodstream and rebuild glycogen reserves. Table 2–4 summarizes information about carbohydrates.

Figure 2–14 The Structure of the Polysaccharide Glycogen. Liver and muscle cells store glucose as the polysaccharide glycogen, a long, branching chain of glucose molecules.



2

Checkpoint

21. **Plant starch and glycogen are both polysaccharides. What monomer do they have in common?**

See the blue Answers tab at the back of the book.

2-11 Lipids often contain a carbon-to-hydrogen ratio of 1:2

Learning Outcome Discuss the structures and functions of lipids.

Like carbohydrates, **lipids** (*lipos*, fat) contain carbon, hydrogen, and oxygen, and the carbon-to-hydrogen ratio is near 1:2. However, lipids contain much less oxygen than do carbohydrates with the same number of carbon atoms.

Table 2–4 Carbohydrates in the Body

Structural Class	Examples	Primary Function	Remarks
Monosaccharides (simple sugars)	Glucose, fructose	Energy source	Manufactured in the body and obtained from food; distributed in body fluids
Disaccharides	Sucrose, lactose, maltose	Energy source	Sucrose is table sugar, lactose is in milk, and maltose is malt sugar found in germinating grain; all must be broken down to monosaccharides before absorption
Polysaccharides	Glycogen	Storage of glucose	Glycogen is in animal cells; other starches and cellulose are within or around plant cells

The hydrogen-to-oxygen ratio is therefore very large. For example, a representative lipid, such as lauric acid (found in coconut, laurel, and palm kernel oils), has a formula of $C_{12}H_{24}O_2$. Lipids may also contain small quantities of phosphorus, nitrogen, or sulfur. Familiar lipids include *fats*, *oils*, and *waxes*. Most lipids are hydrophobic, or insoluble in water, but special transport mechanisms carry them into the bloodstream.

Lipids form essential structural components of all cells. In addition, lipid deposits are important as energy reserves. On average, lipids provide twice as much energy as carbohydrates do, gram for gram, when broken down in the body. When the supply of lipids exceeds the demand for energy, the excess is stored in fat deposits. For this reason, there has been great interest in developing *fat substitutes* that provide less energy, but have the same desirable taste and texture as the fats found in many foods.

Lipids normally make up 12–18 percent of the total body weight of adult men, and 18–24 percent for adult women. Many kinds of lipids exist in the body. We will consider five classes of lipids: *fatty acids*, *eicosanoids*, *glycerides*, *steroids*, and *phospholipids and glycolipids*.

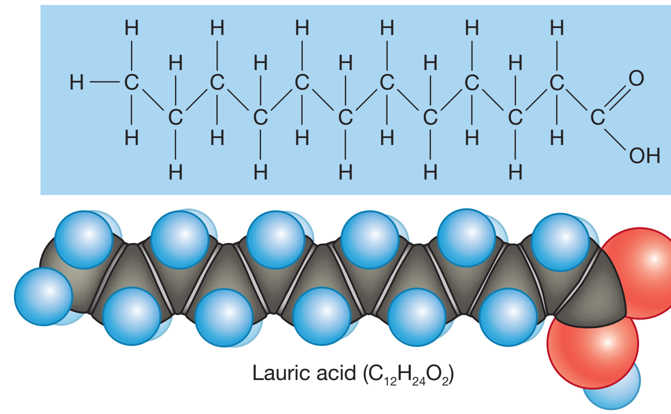
Fatty Acids

Fatty acids are long carbon chains with hydrogen atoms attached. They are one of the monomers of lipids. One end of the carbon chain is always attached to a *carboxyl* (kar-BOK-sil) *group*, —COOH (see **Table 2–3**). The name *carboxyl* should help you remember that a carbon and a hydroxyl (—OH) group are the important structural features of fatty acids. The carbon chain attached to the carboxyl group is known as the *hydrocarbon tail* of the fatty acid. **Figure 2–15a** shows a representative fatty acid, *lauric acid*.

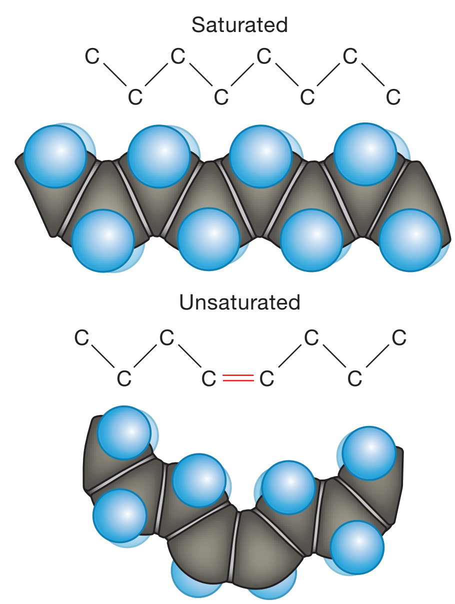
Fatty acids have a very limited solubility in water. When a fatty acid is in solution, only the carboxyl end associates with water molecules, because that is the only hydrophilic portion of the molecule. The hydrocarbon tail is hydrophobic. In general, the longer the hydrocarbon tail, the lower the solubility of the molecule.

Fatty acids may be either saturated or unsaturated (**Figure 2–15b**). These terms refer to the number of hydrogen atoms bound to the carbon atoms in the hydrocarbon tail. In a *saturated* fatty acid, each carbon atom in the tail has four single covalent bonds (see **Figure 2–15a**). Within the tail, two of those bonds bind adjacent carbon atoms, and the other two bind hydrogen atoms. The carbon atom at the end of the tail binds three hydrogen atoms. In an *unsaturated* fatty acid, one or more of the single covalent bonds between the carbon atoms have been replaced by a double covalent bond. As a result, the carbon atoms involved will each bind only one hydrogen atom rather than two. This changes the shape of the hydrocarbon tail, giving it a sharp bend, as you can see in **Figure 2–15b**. The change also affects the way the fatty acid is metabolized.

Figure 2–15 **Fatty Acids.**



a Lauric acid shows two structural characteristics common to all fatty acids: a long chain of carbon atoms and a carboxyl group (—COOH) at one end.



b A fatty acid is either saturated (has single covalent bonds only) or unsaturated (has one or more double covalent bonds). The presence of a double bond causes a sharp bend in the molecule.

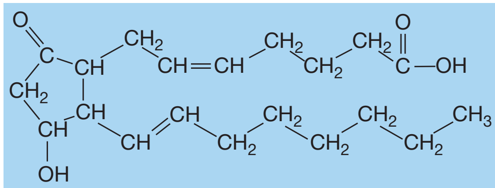
? What type of bond does an unsaturated fatty acid contain that a saturated fatty acid does not?

A *monounsaturated* fatty acid has a single double bond in the hydrocarbon tail. A *polyunsaturated* fatty acid contains two or more double bonds.

Eicosanoids

Eicosanoids (ī-KŌ-sa-noydz) are lipids derived from *arachidonic* (ah-rak-i-DON-ik) *acid*, a fatty acid that must be absorbed in the diet because the body cannot synthesize it. The two

Figure 2–16 Prostaglandins. Prostaglandins contain 20 carbon atoms and a 5-carbon ring.



major classes of eicosanoids are *leukotrienes* and *prostaglandins*. **Leukotrienes** (lū-kō-TRĪ-ēnz) are produced mostly by cells involved with coordinating the responses to injury or disease. We consider leukotrienes in Chapters 18 and 22.

Prostaglandins (pros-tuh-GLAN-dinz) are short-chain fatty acids in which five of the carbon atoms are joined in a ring Figure 2–16). These compounds are released by cells to coordinate or direct local cellular activities, and they are extremely powerful even in small quantities. Virtually all tissues synthesize and respond to them. The effects of prostaglandins vary with their structure and their release site. Prostaglandins released by damaged tissues, for example, stimulate nerve endings and produce the sensation of pain (Chapter 15). Those released in the uterus help trigger the start of labor contractions (Chapter 29).

The body uses several types of chemical messengers. Those that are produced in one part of the body and have effects on distant parts are called *hormones*. Hormones are distributed throughout the body in the bloodstream, but most prostaglandins affect only the area in which they are produced. As a result, prostaglandins are often called *local hormones*. The distinction is not a rigid one, however, as some prostaglandins also enter the bloodstream and affect other areas. We discuss hormones and prostaglandins in Chapter 18.

Glycerides

Unlike monosaccharides, individual fatty acids cannot be strung together in a chain by dehydration synthesis to form a polymer. But they can be attached to a modified simple sugar, **glycerol** (GLIS-er-ol), through a similar reaction. The result is a lipid known as a **glyceride** (GLIS-er-īd). Dehydration synthesis reactions can produce a **monoglyceride** (mon-ō-GLIS-er-īd), consisting of glycerol plus one fatty acid. Subsequent reactions can yield a **diglyceride** (glycerol + two fatty acids) and then a **triglyceride** (glycerol + three fatty acids), as in Figure 2–17. Hydrolysis breaks the glycerides into fatty acids and glycerol (monomers). Comparing Figure 2–17 with Figure 2–13 shows that dehydration synthesis and hydrolysis operate the same way, whether the molecules involved are carbohydrates or lipids.

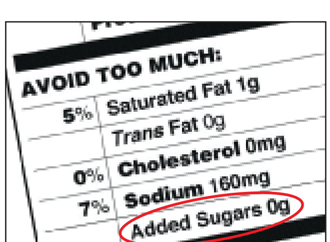
Triglycerides, also known as *triacylglycerols* or *neutral fats*, are important lipid polymers. They have the following important functions:

- *Energy Source*. Fat deposits in the body represent a significant energy reserve. In times of need, the triglycerides are taken apart by hydrolysis, yielding fatty acids that can be broken down to provide energy.

Clinical Note Too Sweet on Sugar?

A baby's first and favorite taste is sweet: mother's milk is rich in *lactose* (milk sugar), a disaccharide of glucose and galactose. This preference for sweet persists throughout life. It is easier to tempt the poor appetite of a frail, elderly person with a bowl of pudding than with a bowl of steamed kale. Manufacturers of processed foods know this.

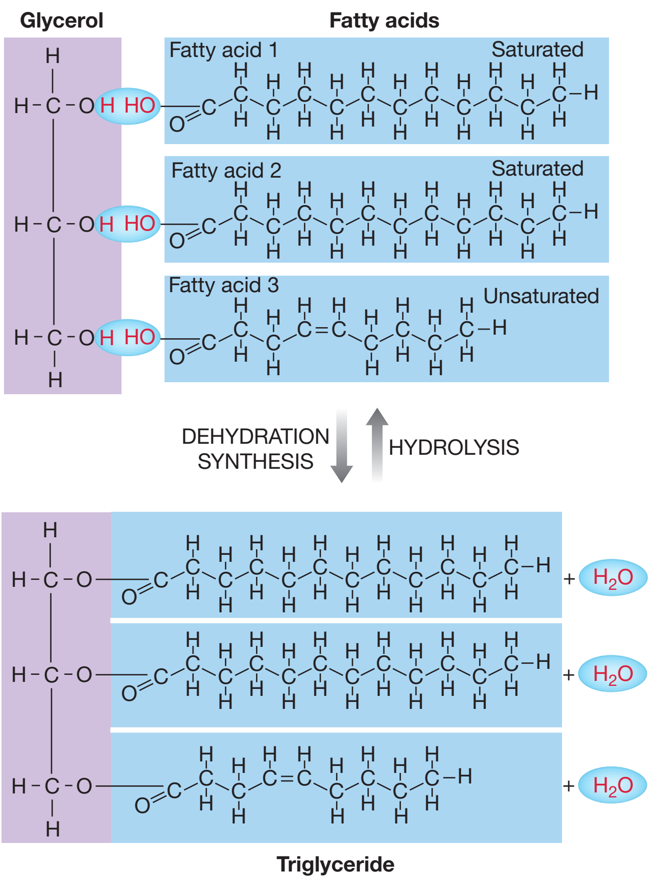
When heart disease became endemic in the United States, holding first place as the killer of Americans, the medical community advocated a low-fat diet for heart health. Artery-clogging fats were removed from manufactured foods—such as cookies, soups, and other boxed, bagged, and frozen products—and replaced with sugar for flavor and mouth appeal.



The sweetening of the American diet has wreaked a new kind of havoc on American health. Dental hygienists see more *dental caries* (cavities). Obesity is climbing at an alarming rate. Grade-school children are developing more *type 2 diabetes*, formerly called "adult-onset diabetes." These serious and potentially fatal diseases generally did not appear until a person had lived several decades with a poor lifestyle.

Glucose is a necessary nutrient. Our body's cells depend on it for fuel; our neurons (brain cells) require it. However, we should meet our glucose needs through complex carbohydrates, or polysaccharides (such as glycogen). In contrast to simple sugars, complex carbohydrates are digested slowly by decomposition reactions in the digestive tract. The component monosaccharides of glucose are released and absorbed gradually, maintaining a steady blood glucose level. In turn, the pancreas is signaled only as needed to make the protein hormone *insulin*, which stimulates the transport of glucose into the body's cells. Complex carbohydrates promote satiety (fullness) and support healthy sugar metabolism.

Figure 2–17 Triglyceride Formation. The formation of a triglyceride involves the attachment of fatty acids to a glycerol molecule through dehydration synthesis. In this example, a triglyceride is formed by the attachment of one unsaturated and two saturated fatty acids to a glycerol molecule.



? What makes fatty acid 3 an unsaturated fatty acid?

- *Insulation*. Fat deposits under the skin serve as insulation, slowing heat loss to the environment. Heat loss across a layer of lipids is only about one-third of the heat loss through other tissues.
- *Protection*. A fat deposit around a delicate organ such as a kidney provides a cushion that protects against bumps or jolts.

Triglycerides are stored in the body as lipid droplets within cells. The droplets absorb and accumulate lipid-soluble vitamins, drugs, or toxins that appear in body fluids. This accumulation has both positive and negative effects. For example, the body's lipid reserves retain both valuable lipid-soluble vitamins (A, D, E, K) and potentially dangerous lipid-soluble pesticides, such as the now-banned DDT.

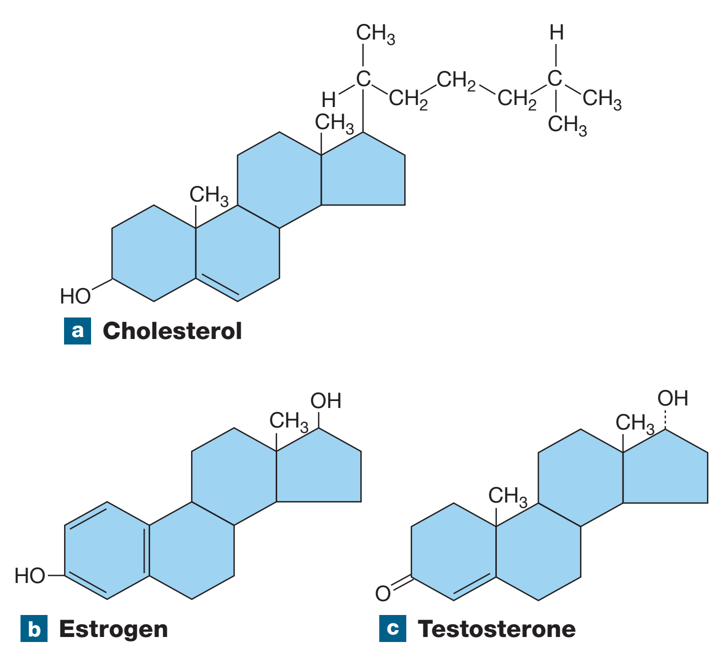
Steroids

Steroids are large lipid molecules that share a distinctive four-ring carbon structure (**Figure 2–18**). They differ in the functional groups that are attached to this basic framework. The steroid **cholesterol** (kō-LES-ter-ol; *chole-*, bile + *stereos*, solid) and related steroids are important for several reasons:

- The outer boundary of all animal cells, called a plasma membrane, contains cholesterol (**Figure 2–18a**). Cells need cholesterol to maintain their plasma membranes, as well as for cell growth and division.
- Steroid hormones are involved in the regulation of sexual function. Examples include the sex hormones *estrogen* and *testosterone* (**Figure 2–18b,c**).
- Steroid hormones are important in the regulation of tissue metabolism and mineral balance. Examples include *corticosteroids* from the adrenal cortex, which play a role in carbohydrate and protein metabolism, and *calcitriol* from the kidneys, a hormone important in the regulation of the body's calcium ion concentrations.
- Steroid derivatives called *bile salts* are required for the normal processing of dietary fats. The liver produces bile salts and secretes them in bile. They interact with lipids in the intestinal tract and assist with the digestion and absorption of lipids.

The body obtains cholesterol in two ways: (1) by absorbing it from animal products in the diet and (2) by synthesizing it. Liver,

Figure 2–18 Steroids Have a Complex Four-Ring Structure. Individual steroids differ in the side chains attached to the carbon rings.



meat, shellfish, and egg yolks are especially rich dietary sources of cholesterol. People with *hypercholesterolemia*, a condition characterized by very high levels of blood cholesterol, have an increased risk of developing a form of heart disease called coronary artery disease (CAD). In CAD, excess cholesterol deposits on arterial walls, forming plaques that obstruct blood flow to the heart. Currently, it is suggested that a healthy person consume no more than 300 mg of cholesterol per day, and others with diabetes, high cholesterol, or heart disease should consume no more than 200 mg per day. Unfortunately, the blood cholesterol level can be difficult to control by dietary restriction alone because the body can synthesize cholesterol as well. In fact, the body makes more than enough, so strict vegetarians do not need to eat animal products to ensure adequate amounts of cholesterol.

Phospholipids and Glycolipids

Phospholipids (FOS-fō-lip-idz) and **glycolipids** (GLĪ-kō-lip-idz) are structurally related, and our cells can synthesize both types of lipids, primarily from fatty acids. In a *phospho*lipid, a *phosphate group* (PO_4^{3-}) links a diglyceride to a nonlipid group (**Figure 2–19a**). There are different types of phospholipids; the one shown in this figure is lecithin. In a *glyco*lipid, a carbohydrate is attached to a diglyceride (**Figure 2–19b**). Note that placing *-lipid* last in these names indicates that the molecule consists primarily of lipid.

The long hydrocarbon tails of phospholipids and glycolipids are hydrophobic, but the opposite ends, the nonlipid *heads*, are hydrophilic. In water, large numbers of these molecules tend to form droplets, or *micelles* (mī-SELZ), with the hydrophilic portions on the outside (**Figure 2–19c**). Most meals contain a mixture of lipids and other organic molecules, and micelles form as the food breaks down in your digestive tract. In addition to phospholipids and glycolipids, micelles may contain other insoluble lipids, such as steroids, glycerides, and long-chain fatty acids.

Phospholipids and glycolipids (as well as cholesterol) are called *structural lipids*, because they help form and maintain intracellular structures called membranes. At the cellular level, *membranes* are sheets or layers composed mainly of lipids. For example, the plasma membrane surrounding each cell is composed primarily of phospholipids. It separates the aqueous solution inside the cell from the aqueous solution outside the cell. Also, various internal membranes subdivide the interior of the cell into specialized compartments, each with a distinctive chemical nature and, as a result, a different function.

The five types of lipids and their characteristics are summarized in **Table 2–5**.

✓ Checkpoint

22. **Describe lipids.**
23. **Which lipids would you find in human plasma membranes?**

See the blue Answers tab at the back of the book.

2-12 Proteins contain carbon, hydrogen, oxygen, and nitrogen and are formed from amino acids

Learning Outcome Discuss the structures and functions of proteins.

Proteins are the most abundant organic molecules in the human body and in many ways the most important. The human body contains many different proteins, and they account for about 20 percent of total body weight. All proteins contain carbon, hydrogen, oxygen, and nitrogen. Smaller quantities of sulfur and phosphorus may also be present. *Amino acids* are simple organic compounds (monomers) that combine to form proteins (polymers). ⮌ p. 45

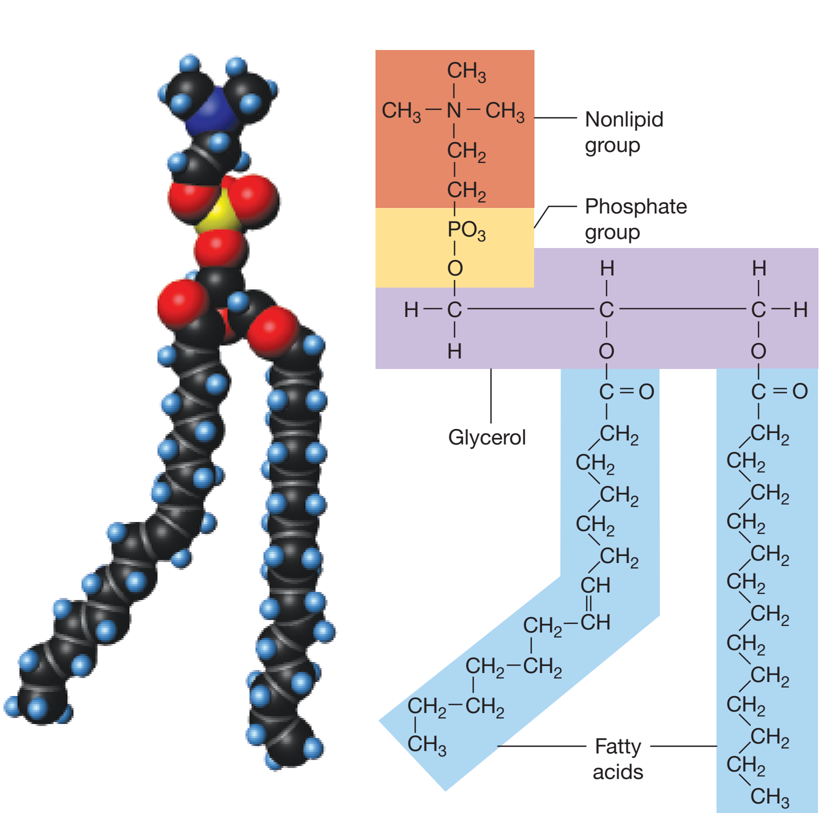
Proteins carry out a variety of essential functions, which we can group into the following major categories:

- *Support. Structural proteins* create a three-dimensional framework for the body. They provide strength, organization, and support for cells, tissues, and organs.

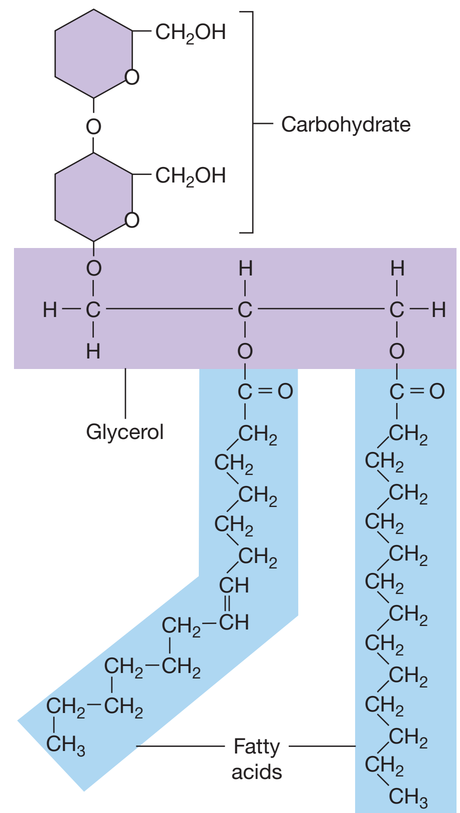
Table 2–5 Representative Lipids and Their Functions in the Body

Lipid Type	Example(s)	Primary Functions	Remarks
Fatty acids	Lauric acid	Energy source	Absorbed from food or synthesized in cells; transported in the blood
Eicosanoids	Prostaglandins, leukotrienes	Chemical messengers coordinating local cellular activities	Prostaglandins are produced in most body tissues
Glycerides	Monoglycerides, diglycerides, triglycerides	Energy source, energy storage, insulation, and physical protection	Stored in fat deposits; must be broken down to fatty acids and glycerol before they can be used as an energy source
Steroids	Cholesterol	Structural component of plasma membranes, hormones, digestive secretions in bile	All have the same four-carbon ring framework
Phospholipids, glycolipids	Lecithin (a phospholipid)	Structural components of plasma membranes	Derived from fatty acids and nonlipid components

Figure 2–19 Phospholipids and Glycolipids.

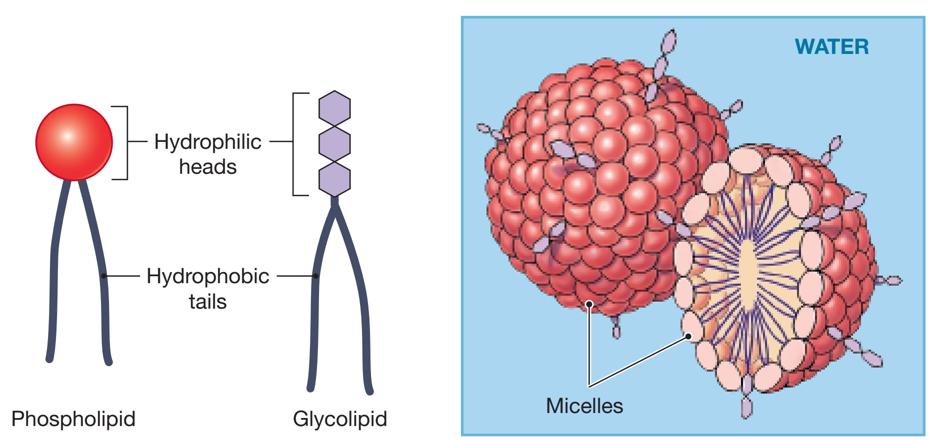


a The phospholipid *lecithin*. In a phospholipid, a phosphate group links a nonlipid molecule to a diglyceride.



b In a glycolipid, a carbohydrate is attached to a diglyceride.

c When large numbers of phospholipids and glycolipids are in water, they form micelles, with the hydrophilic heads facing the water molecules, and the hydrophobic tails on the inside of each droplet.



- *Movement*. *Contractile proteins* bring about muscular contraction. Related proteins are responsible for the movement of individual cells.
- *Transport*. Special *transport proteins* bind many substances for transport in the blood, including insoluble lipids, respiratory gases, special minerals such as iron, and several hormones. These substances would not otherwise be transported in the blood. Other specialized proteins move materials from one part of a cell to another.
- *Buffering*. Proteins provide a buffering action and in this way help prevent dangerous changes in the pH of body fluids.

- *Metabolic Regulation*. Many proteins are enzymes, which as you may recall speed up chemical reactions in cells. The sensitivity of enzymes to environmental factors such as temperature and pH is extremely important in controlling the pace and direction of metabolic reactions.
- *Coordination and Control*. Protein hormones can influence the metabolic activities of every cell in the body or affect the function of specific organs or organ systems.
- *Defense*. Proteins defend the body in many ways. The tough, waterproof proteins of the skin, hair, and nails protect the body from environmental hazards. Proteins called *antibodies* help protect us from disease by taking part in the *immune response*. Special *clotting proteins* restrict bleeding after an injury.

Protein Structure

Proteins are organic polymers that consist of long chains of similar organic molecules called **amino acids**. Twenty different amino acid monomers occur in significant quantities in the body. All 20 amino acids are small, water-soluble molecules. A typical protein contains 1000 amino acids, while the largest protein complexes have 100,000 or more.

Each amino acid consists of five parts (Figure 2–20):

- a central carbon atom,
- a hydrogen atom,
- an *amino group* ($—NH_2$),
- a carboxyl group ($—COOH$),
- an *R group* (a variable *side chain* of one or more atoms that identifies a specific amino acid).

The name *amino acid* refers to the presence of the amino group and the carboxyl group, which all amino acids have in common. At physiological pH levels, the carboxyl group can act as an acid by releasing a hydrogen ion to become a *carboxyl ion* (COO^-). The amino group can act as a base by accepting a hydrogen ion, to become an amino ion ($—NH_3^+$). The result is a molecule that has both positive and negative charges, but a net charge of zero. Such molecules are called *zwitterions*, derived from the German word that means "hybrid."

Figure 2–20 Amino Acids. Each amino acid consists of a central carbon atom to which four different groups are attached: a hydrogen atom, an amino group ($—NH_2$), a carboxyl group ($—COOH$), and a variable side group designated as R.

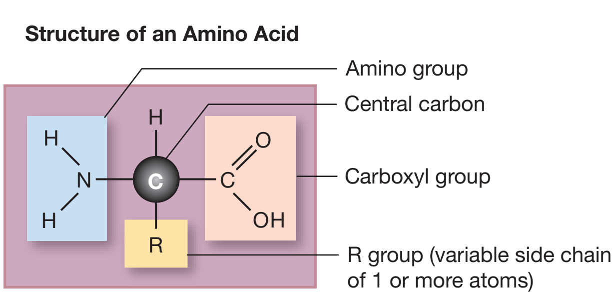
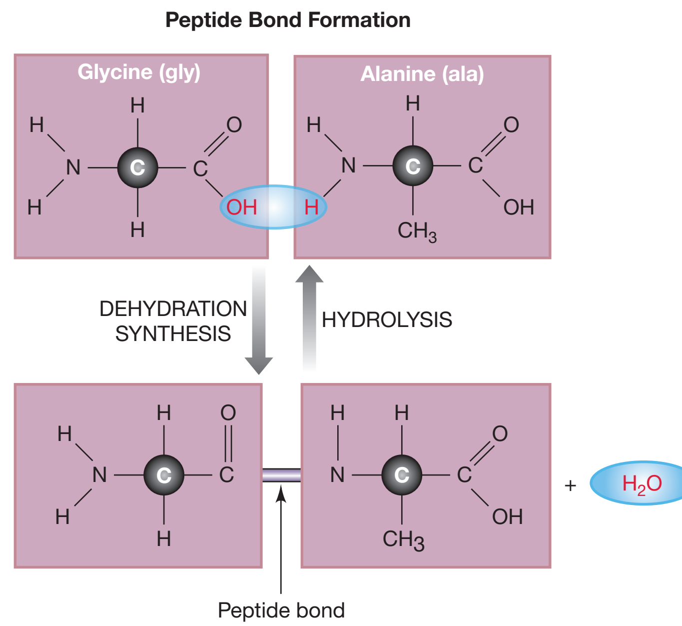


Figure 2–21 The Formation of Peptide Bonds. Peptides form as dehydration synthesis creates a peptide bond between the carboxyl group of one amino acid and the amino group of another. In this example, a peptide bond links the amino acids glycine (for which R = H) and alanine (R = CH_3) to form a dipeptide.



A protein begins to form as amino acids are strung together into long chains. Figure 2–21 shows how dehydration synthesis can link two representative amino acids: *glycine* and *alanine*. This reaction creates a covalent bond between the carboxyl group of one amino acid and the amino group of another. Such a bond is known as a **peptide bond**. Molecules consisting of amino acids held together by peptide bonds are called **peptides**. The molecule created in this example is called a *dipeptide*, because it contains two amino acids.

The chain can be lengthened by the addition of more amino acids. Attaching a third amino acid produces a *tripeptide*. Tripeptides and larger peptide chains are called **polypeptides**. Polypeptides with more than 100 amino acids are usually called proteins. Familiar proteins include *hemoglobin* in red blood cells, *collagen* in skin, bones, and muscles, and *keratin* in fingernails and hair.

The different atoms of the R groups distinguish one amino acid from another, giving each its own chemical properties. For example, different R groups are polar, nonpolar, or electrically charged. Amino acids with nonpolar R groups are hydrophobic, whereas amino acids with polar R groups that form hydrogen bonds with water are hydrophilic. Amino acids with electrically charged R groups are strongly hydrophilic.

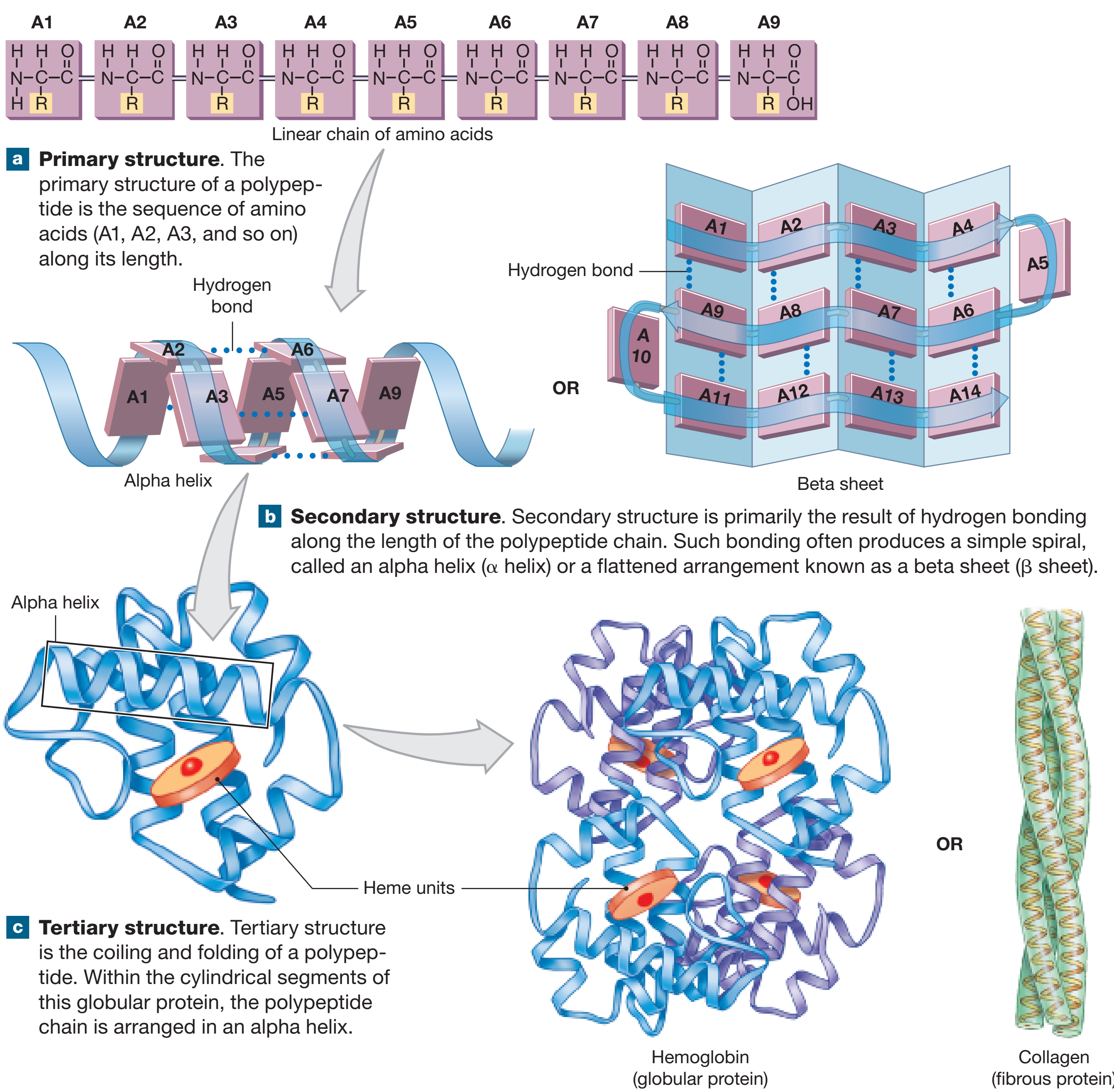
The properties of the R groups contribute to the overall shape and function of proteins.

Protein Shape

The characteristics of a particular protein are determined in part by the R groups on its amino acids. But the properties of a protein are more than just the sum of the properties of its parts, for polypeptides can have highly complex shapes that are important to their function. Proteins can have four levels of structural complexity (Figure 2–22):

1. **Primary structure** is the sequence of amino acids along the length of a single polypeptide (Figure 2–22a).

Figure 2–22 Protein Structure.



a **Primary structure**. The primary structure of a polypeptide is the sequence of amino acids (A1, A2, A3, and so on) along its length.

b **Secondary structure**. Secondary structure is primarily the result of hydrogen bonding along the length of the polypeptide chain. Such bonding often produces a simple spiral, called an alpha helix (α helix) or a flattened arrangement known as a beta sheet (β sheet).

c **Tertiary structure**. Tertiary structure is the coiling and folding of a polypeptide. Within the cylindrical segments of this globular protein, the polypeptide chain is arranged in an alpha helix.

d **Quaternary structure**. Quaternary structure develops when separate polypeptide subunits interact to form a larger molecule. A single hemoglobin molecule contains four globular subunits. Hemoglobin transports oxygen in the blood; the oxygen binds reversibly to the heme units. In collagen, three helical polypeptide subunits intertwine. Collagen is the principal extracellular protein in most organs.

Peptide bonds are responsible for the primary structure of proteins.

2. **Secondary structure** is the shape that results from the presence of hydrogen bonds between atoms at different parts of the polypeptide chain. Hydrogen bonding may create either an *alpha helix* (simple spiral) or *beta sheet* (a flat pleated sheet) (Figure 2–22b). Which one forms depends on where hydrogen bonding takes place between the sequence of amino acids in the polypeptide chain. The alpha helix is the more common form, but a given polypeptide chain may have both helical and pleated sections. Ribbon diagrams are used to represent the three-dimensional structure of proteins. An alpha helix appears as a coiled ribbon, beta sheets as arrows, and less-structured polypeptide chains as narrow ribbons or tubes.
3. **Tertiary structure** is the complex coiling and folding that gives a protein its final three-dimensional shape (Figure 2–22c). Tertiary structure results primarily from hydrophobic and hydrophilic interactions between the R groups of the polypeptide chain and the surrounding water molecules, and to a lesser extent from interactions between the R groups of amino acids in different parts of the molecule. Most such interactions are relatively weak. One, however, is very strong: the *disulfide bond*, a covalent bond that may form between the sulfur atoms of two molecules of the amino acid *cysteine* located at different sites along the chain (not shown). Disulfide bonds create permanent loops or coils in a polypeptide chain.
4. **Quaternary structure** is the interaction between individual polypeptide chains to form a protein complex (Figure 2–22d). Each of the polypeptide subunits has its own secondary and tertiary structures. For example, the protein *hemoglobin* contains four subunits. In Figure 2–22d, the polypeptide subunits with the same color (blue or purple) have the same structure. The hemoglobin in red blood cells binds and transports oxygen. *Collagen*, composed of three windings of alpha helical polypeptides, is the most abundant structural protein and is found in skin, bones, muscles, cartilages, and tendons. Collagen fibers form the framework that supports cells in most tissues. In *keratin*, two alpha helical polypeptides are wound together like the strands of a rope. Keratin is the tough, water-resistant protein at the surface of the skin and in nails and hair.

Fibrous and Globular Proteins

Proteins fall into two general structural classes on the basis of their overall shape and properties:

- **Globular proteins** are compact, generally rounded, and soluble in water. Many enzymes, hormones, and other molecules that circulate in the bloodstream are globular proteins. These proteins can function only if they remain in solution. The unique shape of each globular protein comes from its tertiary structure. Hemoglobin and *myoglobin*, a protein in muscle cells, are both globular proteins. The enzymes that control chemical reactions inside cells are also globular proteins.
- **Fibrous proteins** form extended sheets or strands. Fibrous proteins are tough, durable, and generally insoluble in water. They usually play structural roles in the body. Their shapes are usually due to secondary structure (for proteins with the pleated-sheet form) or quaternary structure (as we just described for collagen and keratin).

Protein Shape and Function

The shape of a protein determines its functional characteristics, and the sequence of amino acids ultimately determines its shape. The 20 amino acids can be linked in an astonishing number of combinations, creating proteins of enormously varied shape and function. Changing only 1 of the 10,000 or more amino acids in a protein can significantly alter the way the protein functions. For example, several cancers and *sickle cell anemia*, a blood disorder, result from changing just a single amino acid in the amino acid sequences of complex proteins.

The tertiary and quaternary shapes of complex proteins depend not only on their amino acid sequence, but also on the local environmental conditions. Small changes in the ionic composition, temperature, or pH of their surroundings can affect the function of proteins. Protein shape can also be affected by hydrogen bonding to other molecules in solution. The significance of these factors is most striking when we consider enzymes, for these proteins are essential to the metabolic operations in every one of our cells.

Enzyme Function

Enzymes are among the most important of all the body's proteins. As noted earlier in this chapter, enzymes catalyze the chemical reactions that sustain life. Almost everything that happens inside the human body does so because a specific enzyme makes it possible.

The reactants in enzymatic reactions are called **substrates**. As in other types of chemical reactions, the interactions among substrates yield specific products. Before an enzyme can function as a catalyst—accelerating a chemical reaction without itself being permanently changed or consumed—the substrates must bind to a special region of the enzyme. This region is called the **active site**. It is typically a groove or pocket into which one or more substrates nestle, like a key fitting into a lock. Weak electrical attractive forces, such as hydrogen bonding, reinforce the physical fit. The tertiary or quaternary structure of the enzyme molecule determines the shape of the active site. Although enzymes are proteins, any organic or inorganic compound that will bind to the active site can be a substrate.

Figure 2–23 A Simplified View of Enzyme Structure and Function. Each enzyme contains a specific active site somewhere on its exposed surface.

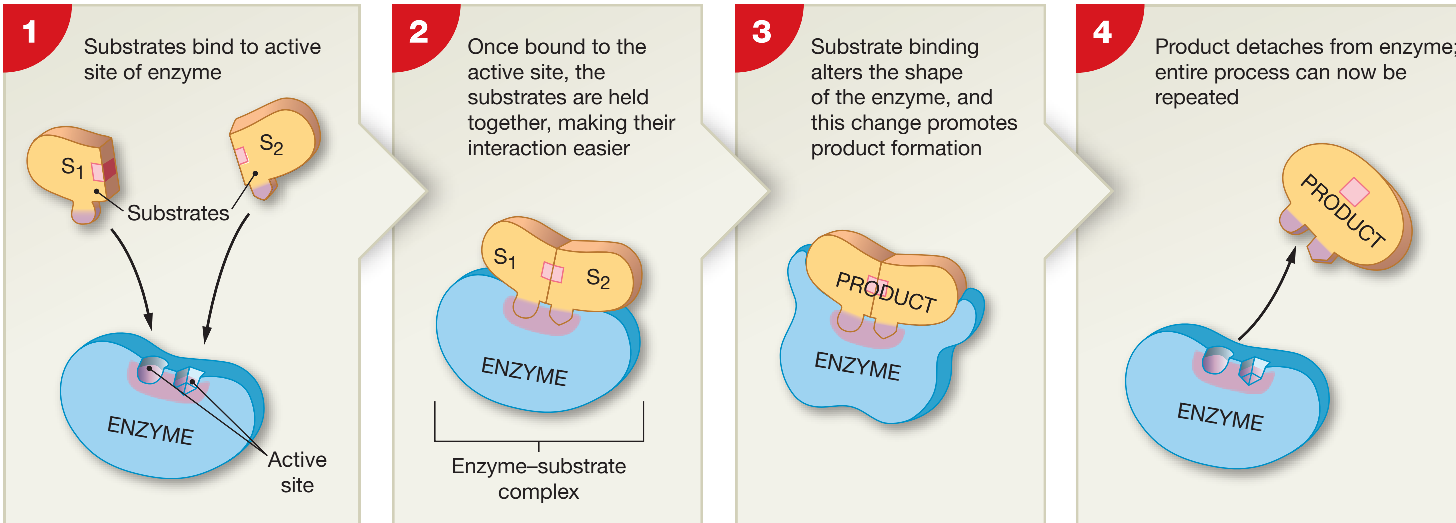


Figure 2–23 presents one example of enzyme structure and function. Substrates bind to the enzyme at its active site (1). Substrate binding produces an *enzyme–substrate complex* (2). Substrate binding typically produces a temporary, reversible change in the shape of the enzyme that may place physical stresses on the substrate molecules, leading to product formation (3). The product is then released, freeing the enzyme to repeat the process (4).

Enzymes work quickly, cycling rapidly between substrates and products. For example, an enzyme providing energy during a muscular contraction performs its reaction sequence 100 times per second. Hydrolytic enzymes can work even faster, breaking down almost 20,000 molecules a second!

Figure 2–23 shows an enzyme that catalyzes a synthesis reaction. Other enzymes may catalyze decomposition reactions, reversible reactions, or exchange reactions. Regardless of the reaction they catalyze, all enzymes share the basic characteristics of specificity, saturation limits, and regulation:

- *Specificity*. Each enzyme catalyzes only one type of reaction, a characteristic called **specificity**. An enzyme's specificity is due to the ability of its active sites to bind only to substrates with particular shapes and charges. For this reason, differences in enzyme structure that do not affect the active site and do not change the response of the enzyme to substrate binding do not affect enzyme function. Such enzyme variants are called *isozymes*.
- *Saturation Limits*. The rate of an enzymatic reaction is directly related to the concentrations of substrate molecules and enzymes. An enzyme molecule must encounter appropriate substrates before it can catalyze a reaction. The higher the substrate concentration, the more frequent these encounters. When substrate concentrations are high enough that every enzyme molecule is cycling through its reaction sequence at top speed, further increases in substrate concentration will not affect the rate of reaction unless additional enzyme molecules are provided. The substrate concentration required to reach the maximum rate of reaction is called the *saturation limit*. An enzyme that has reached its saturation limit is said to be **saturated**. To increase the reaction rate further, the cell must increase the number of enzyme molecules available. This is one important way that cells promote specific reactions.
- *Regulation*. Each cell contains an assortment of enzymes, and any particular enzyme may be active under one set of conditions and inactive under another. Virtually anything that changes the tertiary or quaternary shape of an enzyme can turn it "on" or "off" and in this way control reaction rates inside the cell. Because the shape change is immediate, enzyme activation or inactivation is an important method of short-term control over reaction rates and metabolic pathways. Here we will consider only one example of enzyme regulation: the presence or absence of *cofactors*.

Cofactors and Enzyme Function

A **cofactor** is an ion or a molecule that must bind to an enzyme before substrates can also bind. Without a cofactor, the enzyme is intact but nonfunctional. With the cofactor, the enzyme can catalyze a specific reaction. Examples of ionic cofactors include calcium ion (Ca^{2+}) and magnesium ion (Mg^{2+}), which bind at the enzyme's active site. Cofactors may also bind at other sites, as long as they produce a change in the shape of the active site that makes substrate binding possible.

Coenzymes are nonprotein organic molecules that function as cofactors. Our bodies convert many vitamins into essential coenzymes. *Vitamins*, detailed in Chapter 25, are organic nutrients structurally related to lipids or carbohydrates, but have unique functional roles. Because the human body cannot synthesize most of the vitamins it needs, you must obtain them from your diet.

Temperature and pH Affect Enzyme Function

Each enzyme works best at specific temperatures and pH values. As temperatures rise, protein shape changes and enzyme function deteriorates. Eventually the protein undergoes **denaturation**, a change in tertiary or quaternary structure that makes it nonfunctional. You see permanent denaturation when you fry an egg. As the temperature rises, the proteins in the egg white denature. Eventually, the proteins become completely and irreversibly denatured, forming an insoluble white mass. In the body, death occurs at very high body temperatures (above 43°C, or 110°F) because structural proteins and enzymes soon denature, causing irreparable damage to organs and organ systems.

Enzymes are equally sensitive to changes in pH. *Pepsin*, an enzyme that breaks down food proteins in the stomach, works best at a pH of 2.0 (strongly acidic). Your small intestine contains *trypsin*, another protein-degrading enzyme. Trypsin works only in an alkaline environment, with an optimum pH of 7.7 (weakly basic).

Glycoproteins and Proteoglycans

Glycoproteins (GLĪ-kō-prō-tēnz) and **proteoglycans** (prō-tē-ō-GLĪ-kanz) are combinations of protein and carbohydrate molecules. Glyco*proteins* are large proteins with small carbohydrate groups attached. These molecules may function as enzymes, antibodies, hormones, or protein components of plasma membranes. Glycoproteins in plasma membranes play a major role in identifying normal versus abnormal cells. They are also important in the immune response (Chapter 22). Glycoprotein secretions called *mucins* absorb water to form **mucus**. Mucus coats and lubricates the surfaces of the reproductive and digestive tracts.

Proteo*glycans* are large polysaccharide molecules linked by polypeptide chains. Proteoglycans bind adjacent cells together, and give tissue fluids a viscous (syrupy) consistency.

√ Checkpoint

24. **Describe a protein.**
25. **How does boiling a protein affect its structural and functional properties?**

See the blue Answers tab at the back of the book.

2-13 DNA and RNA are nucleic acids

Learning Outcome Discuss the structures and functions of nucleic acids.

Nucleic (nū -KLĀ-ik) **acids** are large organic molecules composed of carbon, hydrogen, oxygen, nitrogen, and phosphorus. Nucleic acids store and process information at the molecular level inside cells. The two classes of nucleic acid molecules are **deoxyribonucleic** (dē-oks-ē-rī-bō-nū -KLĀ-ik) **acid (DNA)** and **ribonucleic** (rī-bō-nū-KLĀ-ik) **acid (RNA)**. As we will see, these two classes of nucleic acids differ in composition, structure, and function.

The primary role of nucleic acids is to store and transfer information—specifically, information essential to the synthesis of proteins within our cells. The DNA in our cells encodes the information needed to build proteins, while several forms of RNA cooperate to build specific proteins using the information provided by DNA.

DNA contains the instructions for making proteins with correct shapes and therefore correct functions. Those proteins then control our inherited characteristics. For example, by directing the synthesis of structural proteins, DNA determines all of our physical characteristics, including eye color, hair color, and blood type. By directing the synthesis of many functional proteins, including enzymes, DNA regulates all aspects of cellular metabolism, including the creation and destruction of lipids, carbohydrates, and other vital molecules.

In this section, we look at how nucleic acids are structured, and the similarities and differences between DNA and RNA. In Chapter 3 we detail the ways that DNA and RNA work together.

Structure of Nucleic Acids

A nucleic acid consists of one or two long chains of repeating subunits. The individual subunits (monomers) of the chain are called **nucleotides**. Each nucleotide has three parts (**Figure 2–24a**): (1) a pentose (five-carbon sugar) attached to both, (2) a phosphate group and (3) a **nitrogenous** (nitrogen-containing) **base**. The pentose is either **ribose** (in RNA) or **deoxyribose** (in DNA).

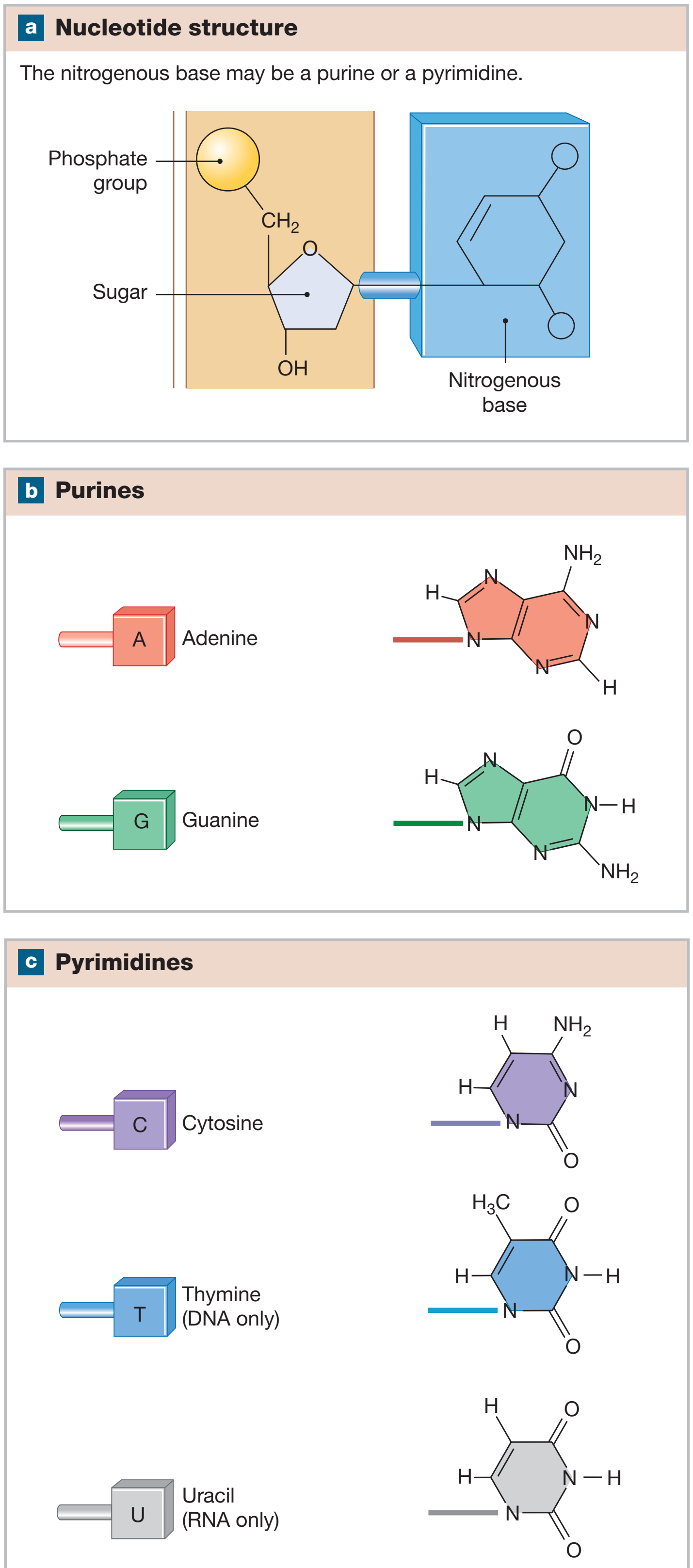
Five nitrogenous bases occur in nucleic acids: **adenine (A)**, **guanine (G)**, **cytosine (C)**, **thymine (T)**, and **uracil (U)**. Adenine and guanine are double-ringed molecules called *purines* (**Figure 2–24b**). The other three bases are single-ringed molecules called *pyrimidines* (**Figure 2–24c**). Both RNA and DNA contain adenine, guanine, and cytosine. Uracil occurs only in RNA and thymine occurs only in DNA.

A nucleotide forms when a phosphate group binds to a pentose already attached to a nitrogenous base. A nucleic acid forms when nucleotides are joined by dehydration synthesis, which attaches the phosphate group of one nucleotide to the sugar of another. The "backbone" of a nucleic acid molecule is a linear sugar-to-phosphate-to-sugar sequence, with the nitrogenous bases projecting to one side.

Comparison of RNA and DNA

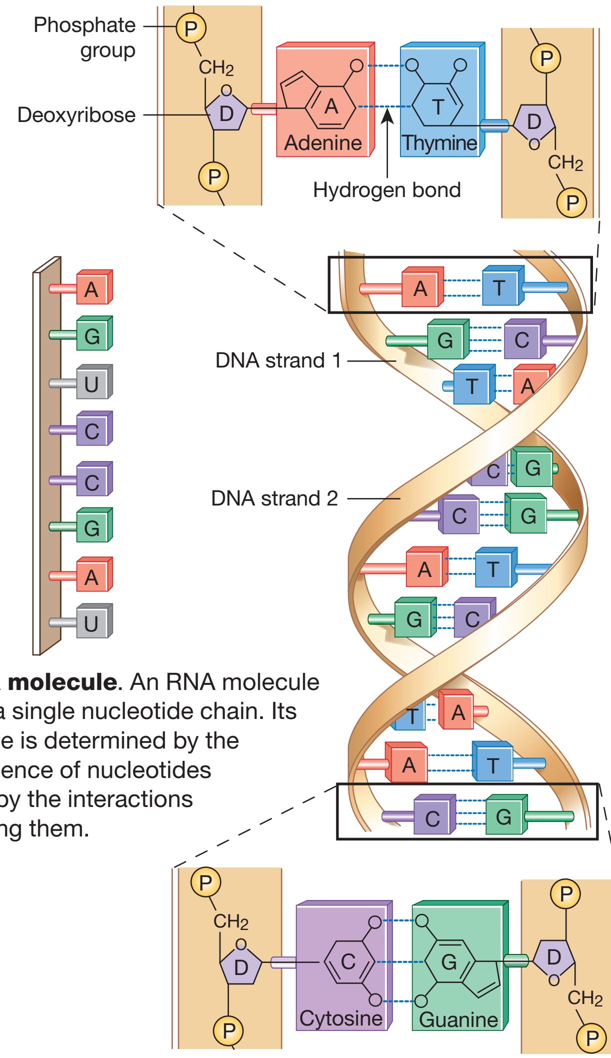
In both DNA and RNA, it is the sequence of nitrogenous bases that carries the information about how to make proteins.

Figure 2–24 Nucleotides and Nitrogenous Bases.



What structural differences make adenine and guanine different from cytosine, thymine, and uracil?

Figure 2–25 The Structure of Nucleic Acids. The nucleic acids RNA and DNA are long chains of nucleotides.



a RNA molecule. An RNA molecule has a single nucleotide chain. Its shape is determined by the sequence of nucleotides and by the interactions among them.

b DNA molecule. A DNA molecule has a pair of nucleotide chains linked by hydrogen bonding between complementary base pairs.

However, there are other important structural differences between the two types of nucleic acids. A molecule of RNA consists of a single chain of nucleotides (**Figure 2–25a**). Its shape depends on the order of the nucleotides and the interactions among them. Our cells have various forms of RNA with different shapes and functions. As we will see in Chapter 3, protein synthesis requires three types of RNA: (1) *messenger RNA (mRNA)*, (2) *transfer RNA (tRNA)*, and (3) *ribosomal RNA (rRNA)*. There are also other types of RNA whose roles are under active research.

A DNA molecule consists of a *pair* of nucleotide chains, with two sugar-phosphate backbones on the outside and the nitrogenous bases projecting inward (**Figure 2–25b**). Hydrogen bonding between opposing nitrogenous bases

Table 2–6 Comparison of RNA and DNA

Characteristic	RNA	DNA
Sugar	Ribose	Deoxyribose
Nitrogenous bases	Adenine (A) Guanine (G) Cytosine (C) Uracil (U)	Adenine Guanine Cytosine Thymine (T)
Number of nucleotides in typical molecule	Varies from fewer than 100 nucleotides to about 50,000	Always more than 45 million
Shape of molecule	Varies with hydrogen bonding along the length of the strand; three main types (mRNA, rRNA, tRNA)	Paired strands coiled in a double helix
Function	Performs protein synthesis as directed by DNA	Stores genetic information that controls protein synthesis

holds the two strands together. The shapes of the nitrogenous bases allow adenine to bond only to thymine, and cytosine to bond only to guanine. As a result, the combinations adenine–thymine (A–T) and cytosine–guanine (C–G) are known as **complementary base pairs**, and the two nucleotide chains of the DNA molecule are known as **complementary strands**. The two strands of DNA twist around one another in a double helix that resembles a spiral staircase. Each step of the staircase corresponds to one complementary base pair.

Through a sequence of events described in Chapter 3 , the cell uses one of the two complementary DNA strands to provide the information needed to synthesize a specific protein. Table 2–6 summarizes our comparison of RNA and DNA.

Checkpoint

26. **Describe a nucleic acid.**
27. **A large organic molecule made of the sugar ribose, nitrogenous bases, and phosphate groups is which kind of nucleic acid?**

See the blue Answers tab at the back of the book.

2-14 ATP is a high-energy compound used by cells

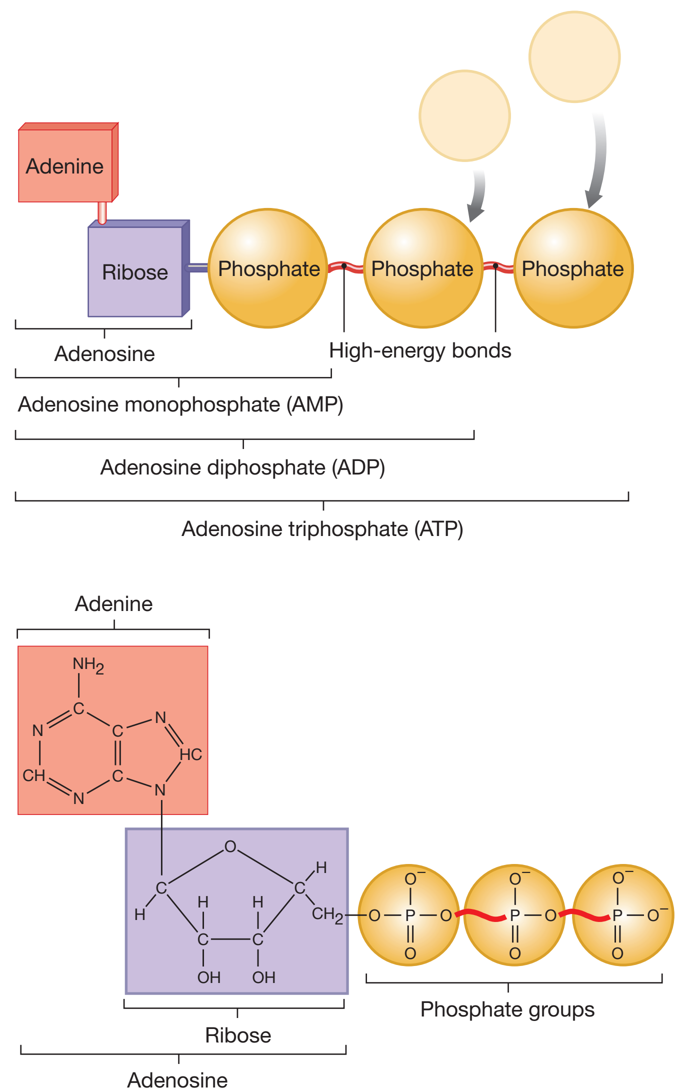
Learning Outcome Discuss the structures and functions of high-energy compounds.

To perform their vital functions, cells must use energy, which they obtain by breaking down organic substrates (catabolism). To be useful, that energy must be transferred from molecule to molecule or from one part of the cell to another.

The usual method of energy transfer involves the creation and breakdown of high-energy bonds by enzymes within cells. A *high-energy bond* is a covalent bond whose breakdown releases energy the cell can use directly. In our cells, a high-energy bond generally binds a phosphate group (PO_4^{3-}) to an organic molecule. The product with such a bond is called a **high-energy compound**. Most high-energy compounds are derived from nucleotides, the building blocks of nucleic acids (Figure 2–26).

The process of attaching a phosphate group to another molecule is called **phosphorylation** (fos-for-i-LĀ-shun). This

Figure 2–26 The Structure of ATP. A molecule of ATP is formed by attaching two phosphate groups to the nucleotide adenosine monophosphate. These two phosphate groups are connected by high-energy bonds incorporating energy released by catabolism. Cells most often obtain quick energy to power cellular operations by removing one phosphate group from ATP, forming ADP (adenosine diphosphate). ADP can later be reconverted to ATP, and the cycle repeated.



? How many phosphorylations does AMP undergo to become ATP?

2

process does not necessarily produce high-energy bonds. For example, in the synthesis of sucrose, a phosphate group is first attached to glucose. The creation of a high-energy compound requires (1) a phosphate group, (2) enzymes capable of catalyzing the reactions involved, and (3) suitable organic substrates to which the phosphate can be added.

The most important such substrate is the nucleotide **adenosine monophosphate (AMP)**, which already contains one phosphate group. Attaching a second phosphate group produces **adenosine diphosphate (ADP)**. A significant energy input is required to convert AMP to ADP, and the second phosphate is attached by a high-energy bond. Even more energy is required to add a third phosphate and create the high-energy compound **adenosine triphosphate (ATP)** (Figure 2–26).

The conversion of ADP to ATP is the most important method of storing energy in our cells. The breakdown of ATP to ADP is the most important method of releasing energy. The relationships involved in this energy transfer can be diagrammed as

$$\text{ADP} + \text{phosphate group} + \text{energy} \rightleftharpoons \text{ATP} + H_2O$$

The hydrolytic breakdown of ATP to ADP requires an enzyme known as **adenosine triphosphatase (ATPase)**, as well as a molecule of water. Throughout life, cells continuously generate ATP from ADP and then use the energy provided by the breakdown of ATP to perform vital functions, such as the synthesis of proteins or the contraction of muscles.

ATP is our most abundant high-energy compound, but there are others. They are typically other nucleotides that have undergone phosphorylation. For example, *guanosine triphosphate (GTP)* is a nucleotide-based high-energy compound that transfers energy in specific enzymatic reactions.

Table 2–7 summarizes the inorganic and organic compounds covered in this chapter.

√ Checkpoint

28. **Describe ATP.**
29. **What molecule is produced by the phosphorylation of ADP?**

See the blue Answers tab at the back of the book.

Table 2–7 Classes of Inorganic and Organic Compounds

Class	Building Blocks (Elements and/or Monomers)	Sources	Functions
INORGANIC			
Water	Hydrogen and oxygen atoms	Absorbed from the diet or generated by metabolism	Solvent; transport medium for dissolved materials and heat; cooling through evaporation; medium for chemical reactions; reactant in hydrolysis
Acids, bases, salts	H^+, OH^-, various anions and cations	Obtained from the diet or generated by metabolism	Structural components; buffers; sources of ions
Dissolved gases	O, C, N, and other atoms	Atmosphere, metabolism	O_2: required for cellular metabolism CO_2: generated by cells as a waste product NO: chemical messenger in cardiovascular, nervous, and lymphatic systems
ORGANIC			
Carbohydrates	C, H, O, in some cases N; CHO in a 1:2:1 ratio Monosaccharide monomers	Obtained from the diet or manufactured in the body	Energy source; some structural role when attached to lipids or proteins; energy storage
Lipids	C, H, O, in some cases N or P; CHO not in 1:2:1 ratio Fatty acids and glycerol monomers	Obtained from the diet or manufactured in the body	Energy source; energy storage; insulation; structural components; chemical messengers; protection
Proteins	C, H, O, N, commonly S Amino acid monomers	20 common amino acids; roughly half can be manufactured in the body, others must be obtained from the diet	Catalysts for metabolic reactions; structural components; movement; transport; buffers; defense; control and coordination of activities
Nucleic acids	C, H, O, N, and P; nucleotides composed of phosphates, sugars, and nitrogenous bases Nucleotide monomers	Obtained from the diet or manufactured in the body	Storage and processing of genetic information
High-energy compounds	Nucleotides joined to phosphates by high-energy bonds	Synthesized by all cells	Storage or transfer of energy

2 Chapter Review

Study Outline

An Introduction to the Chemical Level of Organization p. 28

1. Chemicals combine to form complex structures.

2-1 Atoms are the basic particles of matter p. 28

2. Atoms are the smallest units of matter. They consist of **protons**, **neutrons**, and **electrons**. Protons and neutrons reside in the **nucleus** of an atom.
3. The number of protons in an atom is its **atomic number**. Each **element** includes all the atoms that have the same number of protons and thus the same atomic number.
4. Within an atom, an **electron cloud** surrounds the nucleus. *(Figure 2–1)*
5. The **mass number** of an atom is the total number of protons and neutrons in its nucleus. **Isotopes** are atoms of the same element whose nuclei contain different numbers of neutrons.
6. Electrons occupy an orderly series of **energy levels**, commonly illustrated as **electron shells**. The electrons in the outermost energy level, or **valence shell**, determine an element's chemical properties. *(Figures 2–2, 2–3)*

2-2 Chemical bonds are forces formed by interactions between atoms p. 32

7. Atoms can combine through chemical reactions that create **chemical bonds**. A **molecule** is any chemical structure consisting of atoms held together by covalent bonds. A **compound** is a chemical substance made up of atoms of two or more elements in a fixed proportion.
8. The rules of **chemical notation** are used to describe chemical compounds and reactions. (*Spotlight Figure 2–4*)
9. An **ionic bond** results from the attraction between **ions**, atoms that have gained or lost electrons. **Cations** are positively charged; **anions** are negatively charged. *(Figure 2–5)*
10. Atoms that share electrons to form a molecule are held together by **covalent bonds**. A sharing of one pair of electrons is a **single covalent bond**; a sharing of two pairs is a **double covalent bond**; and a sharing of three pairs is a **triple covalent bond**. A bond with equal sharing of electrons is a **nonpolar covalent bond**; a bond with unequal sharing of electrons is a **polar covalent bond**. *(Figures 2–6, 2–7)*
11. A **hydrogen bond** is a weak, but important, electrical attraction that can affect the shapes and properties of molecules. *(Figure 2–8)*
12. Matter can exist as a *solid*, a *liquid*, or a *gas*, depending on the nature of the interactions among the component atoms or molecules.
13. The **molecular weight** of a molecule or a compound is the sum of the atomic weights of its component atoms.

2-3 Decomposition, synthesis, and exchange reactions are important types of chemical reactions in physiology p. 37

14. A chemical reaction occurs when **reactants** are rearranged to form one or more **products**. Collectively, all the **chemical reactions** in the body constitute its **metabolism**. Through metabolism, cells capture, store, and use energy to maintain homeostasis and to perform essential functions.
15. **Work** is the movement of an object or a change in the physical structure of matter. **Energy** is the capacity to perform work.
16. **Kinetic energy** is the energy of motion. **Potential energy** is stored energy that results from the position or structure of an object. Conversions from potential to kinetic energy (or vice versa) are not 100 percent efficient. Every such energy conversion releases *heat*.
17. A chemical reaction is classified as a **decomposition**, a **synthesis**, or an **exchange reaction**.
18. Cells gain energy to power their functions by **catabolism**, the breakdown of complex molecules. Much of this energy supports **anabolism**, the synthesis of new molecules.
19. All chemical reactions are theoretically reversible. At **equilibrium**, the rates of two opposite reactions are in balance.

2-4 Enzymes speed up reactions by lowering the energy needed to start them p. 39

20. **Activation energy** is the amount of energy required to start a reaction. **Enzymes** are **catalysts**—compounds that accelerate chemical reactions without themselves being permanently changed or consumed. Enzymes promote chemical reactions by lowering the activation energy needed. *(Figure 2–9)*
21. **Exergonic** reactions release energy. **Endergonic** reactions absorb energy.

2-5 Inorganic compounds lack carbon, and organic compounds contain carbon p. 40

22. **Nutrients** are the essential elements and molecules normally obtained from the diet. **Metabolites**, on the other hand, are molecules that can be synthesized or broken down by chemical reactions inside our bodies. Nutrients and metabolites can be broadly categorized as either **inorganic** or **organic compounds**.

2-6 Physiological systems depend on water p. 40

23. Water is the most important constituent of the body.
24. A **solution** is a uniform mixture of two or more substances. It consists of a medium, or **solvent**, in which atoms, ions, or molecules of another substance, or **solute**, are individually dispersed. In *aqueous solutions*, water is the solvent. *(Figure 2–10)*

25. Many inorganic substances, called **electrolytes**, undergo **dissociation**, or **ionization**, in water to form ions. Molecules that interact readily with water molecules are called **hydrophilic**. Those that do not are called **hydrophobic**. *(Figure 2–10; Table 2–2)*

2-7 Body fluid pH is vital for homeostasis p. 43

26. The **pH** of a solution indicates the concentration of hydrogen ions it contains. Solutions are classified as **neutral**, **acidic**, or **basic** *(alkaline)* on the basis of pH. *(Figure 2–11)*

2-8 Acids, bases, and salts have important physiological roles p. 44

27. An **acid** releases hydrogen ions. A **base** removes hydrogen ions from a solution. *Strong acids* and *strong bases* ionize completely. In the case of *weak acids* and *weak bases*, only some of the molecules ionize.
28. A **salt** is an electrolyte whose cation is not a hydrogen ion (OH^+) and whose anion is not a hydroxide ion (OH^-).
29. **Buffers** remove or replace hydrogen ions in solution. Buffers and *buffer systems* in body fluids maintain the pH within normal limits.

2-9 Living things contain organic compounds made up of monomers, polymers, and functional groups p. 45

30. Carbon and hydrogen are the main constituents of **organic compounds**, which generally contain oxygen as well. Identical **monomers** join together through *dehydration synthesis* reactions and form long complex chains called **polymers**. Organic polymers include carbohydrates, lipids, proteins, and nucleic acids.
31. The properties of the different classes of organic monomers and polymers are a result of the presence of *functional groups* of atoms. *(Table 2–3)*

2-10 Carbohydrates contain carbon, hydrogen, and oxygen in a 1:2:1 ratio p. 45

32. **Carbohydrates** are most important as an energy source for metabolic processes. The three major types of carbohydrates are **monosaccharides** *(simple sugars)*, **disaccharides**, and **polysaccharides**. Disaccharides and polysaccharides form from monosaccharide monomers by **dehydration synthesis**. *(Figures 2–12 to 2–14; Table 2–4)*

2-11 Lipids often contain a carbon-to-hydrogen ratio of 1:2 p. 47

33. **Lipids** include *fats*, *oils*, and *waxes*. Most are insoluble in water. The five important classes of lipids are **fatty acids**, **eicosanoids**, **glycerides**, **steroids**, and **phospholipids** and **glycolipids**. *(Figures 2–15 to 2–19; Table 2–5)*
34. **Triglycerides** *(neutral fats)* consist of three fatty acid molecules attached by dehydration synthesis to a molecule of **glycerol**. **Diglycerides** consist of two fatty acids and glycerol. **Monoglycerides** consist of one fatty acid plus glycerol. Fatty acids and glycerol are lipid monomers. *(Figure 2–17)*
35. Steroids (1) are components of plasma membranes, (2) include sex hormones and hormones regulating metabolic activities, and (3) are important in lipid digestion. *(Figure 2–18)*
36. **Phospholipids** and **glycolipids** are structural lipids that are components of *micelles* and plasma membranes *(Figure 2–19)*.

2-12 Proteins contain carbon, hydrogen, oxygen, and nitrogen and are formed from amino acids p. 51

37. **Proteins** perform a variety of essential functions in the body. Seven important types of proteins are *structural proteins*, *contractile proteins*, *transport proteins*, *buffering proteins*, *enzymes*, *hormones*, and *antibodies*.
38. Proteins are organic polymers made up of chains of **amino acids**. Each amino acid monomer consists of an *amino group*, a *carboxyl group*, a *hydrogen atom*, and an *R group (side chain)* attached to a central carbon atom. A **polypeptide** is a linear sequence of amino acids held together by **peptide bonds**; **proteins** are polypeptides containing over 100 amino acids. *(Figures 2–20, 2–21)*
39. The four levels of protein structure are **primary structure** (amino acid sequence), **secondary structure** (amino acid interactions, such as hydrogen bonds), **tertiary structure** (complex folding, disulfide bonds, and interaction with water molecules), and **quaternary structure** (formation of protein complexes from individual subunits). **Globular proteins**, such as *myoglobin* and *hemoglobin*, are generally rounded and water soluble. **Fibrous proteins**, such as *collagen* and *keratin*, are elongated, tough, durable, and generally insoluble. *(Figure 2–22)*
40. The reactants in an enzymatic reaction, called **substrates**, interact to yield a product by binding to the enzyme's **active site**. **Cofactors** are ions or molecules that must bind to the enzyme before the substrates can bind. **Coenzymes** are organic cofactors commonly derived from *vitamins*. *(Figure 2–23)*
41. The shape of a protein determines its functional characteristics. Each protein works best at an optimal combination of temperature and pH and will undergo temporary or permanent **denaturation**, or change in shape, at temperatures or pH values outside the normal range.

2-13 DNA and RNA are nucleic acids p. 57

42. **Nucleic acids** store and process information at the molecular level. The two kinds of nucleic acids are **deoxyribonucleic acid (DNA)** and **ribonucleic acid (RNA)**. *(Figures 2–24, 2–25; Table 2–6)*
43. Nucleic acids are organic polymers made up of chains of **nucleotides**. Each nucleotide contains a sugar, a phosphate group, and a **nitrogenous base**. The sugar is *ribose* in RNA and *deoxyribose* in DNA. DNA is a two-stranded double helix containing the nitrogenous bases **adenine**, **guanine**, **cytosine**, and **thymine**. RNA consists of a single strand and contains **uracil** instead of thymine.

2-14 ATP is a high-energy compound used by cells p. 59

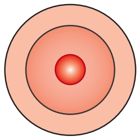
44. Cells store energy in the *high-energy bonds* of **high-energy compounds**. The most important high-energy compound is **ATP (adenosine triphosphate)**. Cells make ATP by adding a phosphate group to **ADP (adenosine diphosphate)** through **phosphorylation**. When ATP is broken down to ADP and phosphate, energy is released. Cells can use this energy to power essential activities. *(Figure 2–26; Table 2–7)*

Review Questions

See the blue Answers tab at the back of the book.

LEVEL 1 Reviewing Facts and Terms

1. An oxygen atom has eight protons. **(a)** Sketch in the arrangement of electrons around the nucleus of the oxygen atom in the following diagram. **(b)** How many more electrons will it take to fill the outermost energy level?



Oxygen atom

2. What is the following type of decomposition reaction called?

   $ABCD + H_2O \longrightarrow ABCH + DOH$
3. The subatomic particle with the least mass **(a)** carries a negative charge, **(b)** carries a positive charge, **(c)** plays no part in the atom's chemical reactions, **(d)** is found only in the nucleus.
4. Isotopes of an element differ from each other in the number of **(a)** protons in the nucleus, **(b)** neutrons in the nucleus, **(c)** electrons in the outer shells, **(d)** a, b, and c are all correct.
5. The number and arrangement of electrons in an atom's outer energy level (valence shell) determine the atom's **(a)** atomic weight, **(b)** atomic number, **(c)** molecular weight, **(d)** chemical properties.
6. All organic compounds in the human body contain all of the following elements *except* **(a)** hydrogen, **(b)** oxygen, **(c)** carbon, **(d)** calcium, **(e)** both a and d.
7. A substance containing atoms of different elements that are bonded together is called a(n) **(a)** molecule, **(b)** compound, **(c)** mixture, **(d)** isotope, **(e)** solution.
8. All the chemical reactions that occur in the human body are collectively referred to as **(a)** anabolism, **(b)** catabolism, **(c)** metabolism, **(d)** homeostasis.
9. Which of the following chemical equations illustrates a typical decomposition reaction?

   (a) $A + B \longrightarrow AB$

   (b) $AB + CD \longrightarrow AD + CB$

   (c) $2A_2 + B_2 \longrightarrow 2A_2B$

   (d) $AB \longrightarrow A + B$
10. The speed, or rate, of a chemical reaction is influenced by **(a)** the presence of catalysts, **(b)** the temperature, **(c)** the concentration of the reactants, **(d)** a, b, and c are all correct.
11. A pH of 7.8 in the human body typifies a condition referred to as **(a)** acidosis, **(b)** alkalosis, **(c)** dehydration, **(d)** homeostasis.
12. A(n) _____ is a solute that dissociates to release hydrogen ions, and a(n) _____ is a solute that removes hydrogen ions from solution. **(a)** base, acid, **(b)** salt, base, **(c)** acid, salt, **(d)** acid, base.
13. Special catalytic molecules called _____ speed up chemical reactions in the human body. **(a)** enzymes, **(b)** cytozymes, **(c)** cofactors, **(d)** activators, **(e)** cytochromes.
14. Which of the following is *not* a function of a protein? **(a)** support, **(b)** transport, **(c)** metabolic regulation, **(d)** storage of genetic information, **(e)** movement.
15. Complementary base pairing in DNA includes the pairs **(a)** adenine–uracil and cytosine–guanine, **(b)** adenine–thymine and cytosine–guanine, **(c)** adenine–guanine and cytosine–thymine, **(d)** guanine–uracil and cytosine–thymine.
16. What are the three subatomic particles in atoms?
17. What four major classes of organic compounds (polymers) are found in the body?
18. List three important functions of triglycerides (neutral fats) in the body.
19. List seven major functions performed by proteins.
20. **(a)** What three basic components make up a nucleotide of DNA? **(b)** What three basic components make up a nucleotide of RNA?
21. What three components are required to create the high-energy compound ATP?

LEVEL 2 Reviewing Concepts

22. If a polypeptide contains 10 peptide bonds, how many amino acids does it contain? **(a)** 9, **(b)** 10, **(c)** 11, **(d)** 12.
23. A dehydration synthesis reaction between glycerol and a single fatty acid would yield a(n) **(a)** micelle, **(b)** omega-3 fatty acid, **(c)** triglyceride, **(d)** monoglyceride, **(e)** diglyceride.
24. Explain how enzymes function in chemical reactions.
25. What is a salt? How does a salt differ from an acid or a base?
26. Explain the differences among nonpolar covalent bonds, polar covalent bonds, and ionic bonds.
27. In an exergonic reaction, **(a)** large molecules are broken down into smaller ones, **(b)** small molecules are assembled into larger ones, **(c)** molecules are rearranged to form new molecules, **(d)** molecules move from reactants to products and back, **(e)** energy is released during the reaction.
28. The hydrogen bonding that occurs in water is responsible for all of the following *except* **(a)** the high boiling point of water, **(b)** the low freezing point of water, **(c)** the ability of water to dissolve nonpolar substances, **(d)** the ability of water to dissolve inorganic salts, **(e)** the surface tension of water.
29. A sample that contains an organic molecule has the following constituents: carbon, hydrogen, oxygen, nitrogen, and phosphorus. Is the molecule more likely to be a carbohydrate, a lipid, a protein, or a nucleic acid?

LEVEL 3 Critical Thinking and Clinical Applications

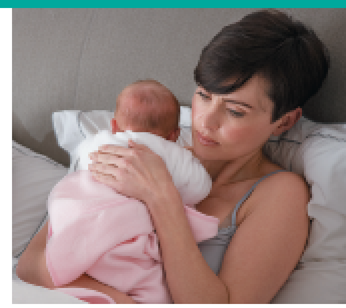
30. An atom of the element calcium has 20 protons and 20 neutrons. Determine the following information about calcium: **(a)** number of electrons, **(b)** atomic number, **(c)** atomic weight, **(d)** number of electrons in each energy level.
31. A certain reaction pathway consists of four steps. How would decreasing the amount of enzyme that catalyzes the second step affect the amount of product produced at the end of the pathway?
32. An important buffer system in the human body involves carbon dioxide (CO_2) and bicarbonate ion (HCO_3^-) in the reversible reaction

$$CO_2 + H_2O \rightleftharpoons H_2CO_3 \rightleftharpoons H^+ + HCO_3^-$$

If a person becomes excited and exhales large amounts of CO_2, how will the pH of the person's body be affected?

2

CLINICAL CASE Wrap-Up What Is Wrong with My Baby?



Baby Sean has *cystic fibrosis (CF),* a life-threatening genetic disease. He inherited a defective gene from each parent. This faulty gene adversely affects the transport of salt—sodium chloride (NaCl)—into and out of cells. As a result, thick, sticky secretions are produced in both the digestive and respiratory systems.

In someone with CF, the digestive juices produced are so thick they clog in the pancreas and cannot get into the small intestine. Digestive juices contain enzymes that break down carbohydrates, lipids, and proteins in food so that they can be absorbed and used by cells. Without these enzymes, Sean's food passes right through him without being digested, leading to weight loss and fatty stools. In addition, the abnormal secretions clog the lungs, making Sean short of breath, wheezy, and susceptible to repeated lung infections. Finally, CF makes his skin taste very salty because of salt being lost in the sweat.

The diagnosis of cystic fibrosis is made with a sweat test. A sweat-producing chemical is applied to an area of Sean's skin, and the sweat is collected and tested for chloride concentration.

Sean's treatment starts immediately. His parents give him digestive enzymes before each feeding so that he can absorb nutrients. Sean's parents learn chest physical therapy: For 20–60 minutes, twice per day, they percuss (clap forcefully with a cupped hand) and vibrate (shake with an open hand) his chest to loosen the thick mucus so he can cough it up and out. Soon Sean gains weight and begins to thrive, but this therapy will be lifelong.

1. What undigested substances would have made Sean's stools greasy and foamy?
2. Why are digestive enzymes necessary for life?

See the blue Answers tab at the back of the book.

Related Clinical Terms

artificial sweetener: Organic molecules that can stimulate taste buds and provide a sweet taste to foods without adding substantial amounts of calories to the diet.

heavy metal: The term used for a group of elements on the "heavier" end of the periodic table of elements. Some heavy metals—cobalt, copper, iron, manganese, molybdenum, vanadium, strontium, and zinc—are essential to health in trace amounts. Others are nonessential and can be harmful to health in excessive amounts. These include cadmium, antimony, chromium, mercury, lead, and arsenic.