The Chemical Level of Organization

2

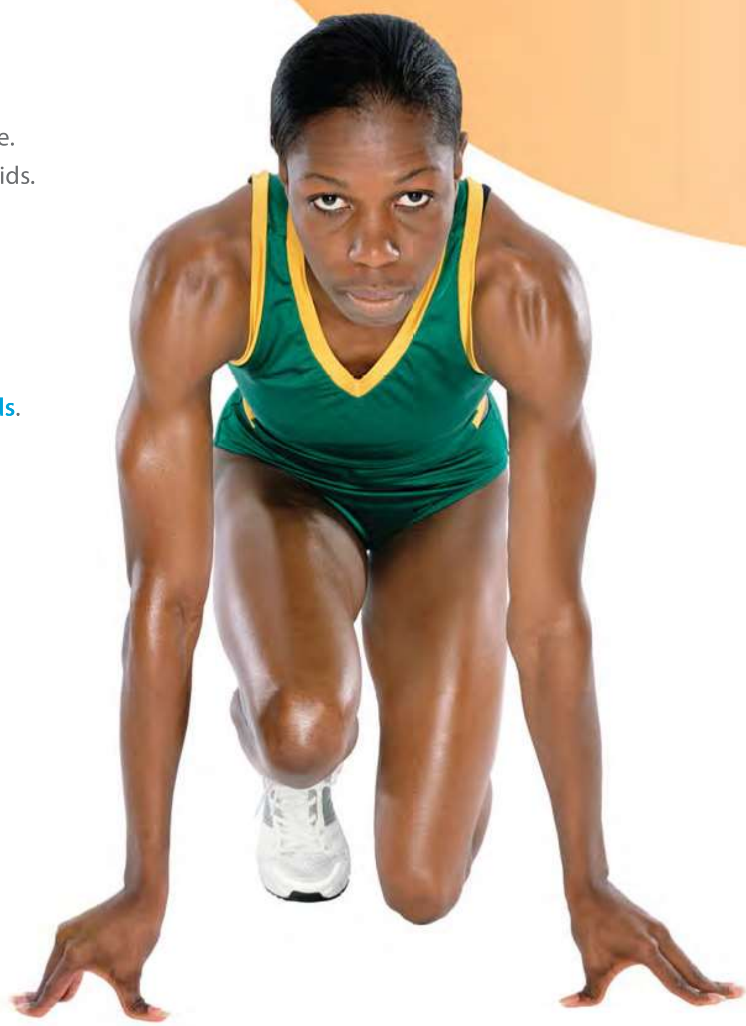
Learning Outcomes

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

2-1 Describe an **atom** and how **atomic structure** affects interactions between atoms.
2-2 Compare the ways in which atoms combine to form **molecules and compounds**.
2-3 Distinguish among the major types of **chemical reactions** that are important for studying physiology.
2-4 Describe the crucial role of **enzymes** in metabolism.
2-5 Distinguish between **organic and inorganic compounds**.
2-6 Explain how the chemical **properties of water** make life possible.
2-7 Discuss the importance of **pH** and the role of buffers in body fluids.
2-8 Describe the physiological roles of **inorganic compounds**.
2-9 Discuss the structures and functions of **carbohydrates**.
2-10 Discuss the structures and functions of **lipids**.
2-11 Discuss the structures and functions of **proteins**.
2-12 Discuss the structures and functions of **nucleic acids**.
2-13 Discuss the structures and functions of **high-energy compounds**.
2-14 Explain the relationship between **chemicals and cells**.

Clinical Notes

Spotlight



An Introduction to the Chemical Level of Organization

This chapter considers the structure of atoms, the basic chemical building blocks. You will also learn how atoms can be combined to form increasingly complex structures.

2-1 Atoms are the basic particles of matter

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.** Although many different subatomic particles exist, only three are important for understanding the chemical properties of matter: *protons*, *neutrons*, and *electrons*. Protons and neutrons are similar in size and mass, but **protons** (p^+) have a positive electrical charge, whereas **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. The mass of an atom is therefore 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 the component atoms.

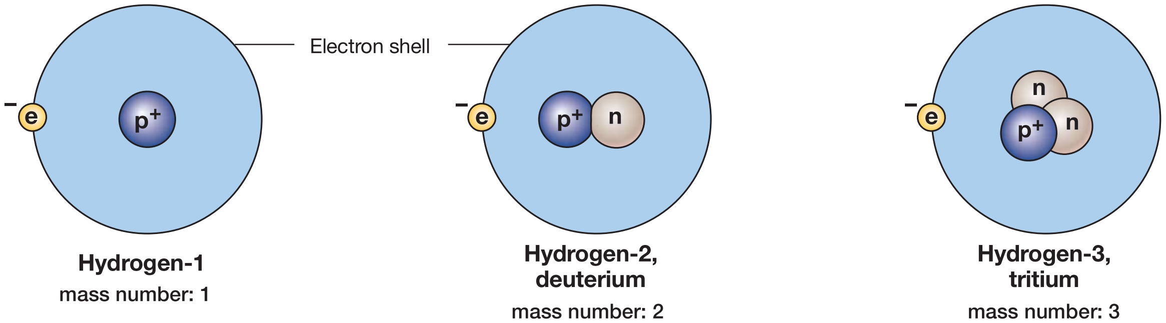
Atomic Structure

Atoms normally contain equal numbers of protons and electrons. The number of protons in an atom is known as its **atomic number.** Hydrogen (H) is the simplest atom, with an atomic number of 1. Thus, an atom of hydrogen contains one proton, and one electron as well. 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.** We often illustrate atomic structure in the simplified form shown in **Figure 2–1a**. 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 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 most conveniently reported in nanometers (NAN-ō-mē-ter) (nm). One nanometer is 10^{-9} meter (0.000000001 m). The very largest atoms approach 0.5 nm in diameter (0.0000000005 m, or 0.00000002 in.).

Figure 2–1 The Structure of Hydrogen Atoms. Three forms of hydrogen atoms are shown using the two-dimensional electron-shell model, which indicates the electron cloud surrounding the nucleus.



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

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

c A tritium (^3H) nucleus contains a pair of neutrons in addition to the proton.

2

Elements and Isotopes

An **element** is a pure substance composed of atoms of only one kind; because atoms are the smallest particles of an element that still retain the characteristics of that element, each element has uniform composition and properties. Each element includes all the atoms that have the same number of protons, and thus the same atomic number. Only 92 elements exist in nature, although about two dozen additional elements have been created through nuclear reactions in research laboratories. Every element has a chemical symbol, an abbreviation recognized by scientists everywhere. Most of the symbols are easily connected 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*.

Because atomic nuclei are unaltered by ordinary chemical processes, elements cannot be changed or broken down into simpler substances in chemical reactions. Thus, an atom of carbon always remains an atom of carbon, regardless of the chemical events in which it may take part.

The human body consists of many elements, and the 13 most abundant elements are shown in **Table 2–1**. The human body also contains atoms of another 14 elements—called *trace elements*—that are present in very small amounts.

The atoms of a single element can differ in the number of neutrons in the nucleus. For example, although most hydrogen nuclei consist of a single proton, 0.015 percent also contain one neutron, and a very small percentage contain two neutrons (**Figure 2–1**). Atoms of the same element whose nuclei contain the same number of protons, but different numbers of neutrons, are called **isotopes.** Different isotopes of an element have essentially identical chemical properties, and are alike except on the basis of mass. The **mass number**—the total number of protons plus neutrons in the nucleus—is used to designate isotopes. Thus, the three isotopes of hydrogen are hydrogen-1, or ^{1}H, with one proton and one electron (**Figure 2–1a**); hydrogen-2, or ^{2}H, also known as *deuterium*, with one proton, one electron, and one neutron (**Figure 2–1b**); and hydrogen-3, or ^{3}H, also known as *tritium*, with one proton, one electron, and two neutrons (**Figure 2–1c**).

The nuclei of some isotopes are unstable and spontaneously break down and emit subatomic particles or radiation in measurable amounts. Such isotopes are called **radioisotopes,** and the breakdown process is called *radioactive decay*. Strongly radioactive isotopes are dangerous, because the emissions can break molecules apart, destroy cells, and otherwise damage living tissues. Weakly radioactive isotopes are sometimes used in diagnostic procedures to monitor the structural or functional characteristics of internal organs.

Radioisotopes differ in how rapidly they decay. The decay rate of a radioisotope is commonly expressed in terms of its **half-life:** the time required for half of a given amount of the isotope to decay. The half-lives of radioisotopes range from fractions of a second to billions of years.

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

Atomic Weights

A typical *oxygen* atom, which has an atomic number of 8, contains eight protons and eight 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 is known as its **atomic weight.**

The unit used to express atomic weight is the *atomic mass unit* (amu). One atomic mass unit is very close to the weight of a single proton or neutron. Thus, the atomic weight 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 mass number that reflects the proportions of different isotopes. That is, the atomic weight of an element is usually very close to the mass number of the most common isotope of that element. For example, the atomic number of hydrogen is 1, but the atomic weight of hydrogen is 1.0079, primarily because some hydrogen atoms (0.015 percent) have a mass number of 2, and even fewer have a mass number of 3.

Atoms participate in chemical reactions in fixed numerical ratios. To form water, for example, exactly two atoms of hydrogen combine with one atom of oxygen. But individual atoms are far too small and too numerous to be counted, so chemists use a unit called the *mole*. For any element, a **mole** (abbreviated *mol*) is a quantity with a weight in grams equal to that element's atomic weight. The mole is useful because one mole of a given element always contains the same number of atoms as one mole of any other element. 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 they 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; 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. Although the electrons in an energy level may travel in complex patterns around the nucleus, for our purposes the patterns can be diagrammed as a series of concentric electron shells. The first electron shell (the one closest to the nucleus) corresponds to the lowest energy level.

Each energy level is limited in the number of electrons it can hold. The first energy level can hold at most 2 electrons, and for our purposes, the next two levels can each hold up to 8 electrons. 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, respectively, of the periodic table. The electrons in an atom occupy successive shells in an orderly manner: The first energy level is filled before any electrons enter the second, and the second energy level is filled before any electrons enter the third.

The outermost energy level forms the "surface" of the atom and is called the **valence shell.** The number of electrons in this level determines the chemical properties of the element. Atoms with unfilled energy levels are unstable—that is, they will react with other atoms, usually in ways that result in full outer electron

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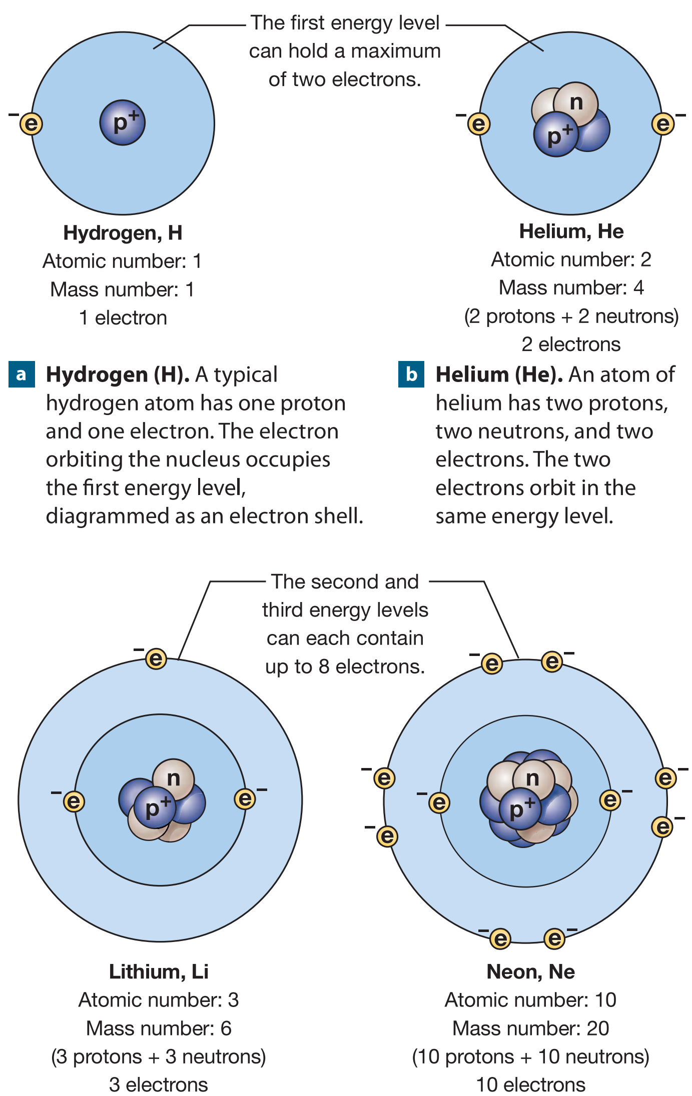
Figure 2–2 The Arrangement of Electrons into Energy Levels.



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

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

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

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

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shells. In contrast, atoms with a filled outermost energy level are stable and therefore do not readily react with other atoms.

As indicated in **Figure 2–2a**, a hydrogen atom has one electron in the first energy level, and that level is thus unfilled. A hydrogen atom readily reacts with other atoms. A helium atom has 2 electrons in its first energy level (**Figure 2–2b**). Because its outer energy level is filled, a helium atom is very stable; it will not ordinarily react with other atoms. A lithium atom has 3 electrons (**Figure 2–2c**). Its first energy level can hold only 2 of them, so lithium has a single electron in a second, unfilled energy level. Like hydrogen, lithium is also unstable and reactive. The second energy level is filled in a neon atom, which has an atomic number of 10 (**Figure 2–2d**). Neon atoms, like helium atoms and other elements in the last column of the periodic table, are thus very stable. The atoms that are most important to biological systems are unstable, because those atoms interact to form larger structures (**Table 2–1**, p. 28).

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 atom interactions

Elements that do not readily participate in chemical processes are said to be *inert*. The noble gases, helium, neon, and argon have filled outermost energy levels. These elements are also called *inert gases*, because their atoms neither react with one another nor combine with atoms of other elements. Elements with unfilled outermost energy levels, such as hydrogen and lithium, are called *reactive*, because they readily interact or combine with other atoms. Reactive atoms achieve stability by gaining, losing, or sharing electrons to fill their outermost energy level. The interactions often involve the formation of **chemical bonds,** which hold the participating atoms together once the reaction has ended. In the sections that follow, we will consider three basic types of chemical bonds: *ionic bonds*, *covalent bonds*, and *hydrogen bonds*.

When chemical bonding occurs, the result is the creation of new chemical entities called *molecules* and *compounds*. The term **molecule** refers to any chemical structure consisting of atoms held together by covalent bonds. A **compound** is a pure chemical substance made up of atoms of two or more different elements, regardless of the type of bond joining them. The two categories overlap, but they aren't the same. Not all molecules are compounds, because some molecules consist of atoms of only one element. (Two oxygen atoms, for example, can be joined by a covalent bond 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 ionic bonds rather than covalent bonds. Many substances, however, belong to both categories. Water is a compound because it contains two different elements—hydrogen and oxygen—and it consists of molecules, because the hydrogen and oxygen atoms are held together by covalent bonds. As we will see in subsequent sections, most biologically important compounds, from carbohydrates to DNA, are molecular.

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 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 very useful rules of this system are listed in **Spotlight Figure 2–7.**

Ionic Bonds

As the name implies, ionic bonds form between ions. **Ions** are atoms or molecules that carry 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). **Ionic bonds** are chemical bonds created by the electrical attraction between anions and cations.

Tips & Tricks

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.

Ions have an unequal number of protons and electrons. Atoms become ions by losing or gaining electrons. We assign a value of +1 to the charge on a proton; the charge on an electron is −1. When the number of protons is equal to the number of electrons, an atom is electrically neutral. An atom that loses an electron becomes a cation with a charge of +1, because it then has one 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.

In the formation of an ionic bond,

- 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.
- attraction between the opposite charges then draws the two ions together.

The formation of an ionic bond is illustrated in **Figure 2–3a**. 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, their presence has no effect on the formation of ions or ionic bonds.) Electrons fill the first and second energy levels, and a single electron occupies the outermost level. Losing that 1 electron would give the sodium atom a full outermost energy level—the second 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 away the electron: The electron must be donated to an electron acceptor. A chlorine atom has 7 electrons in its outermost energy level, 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.

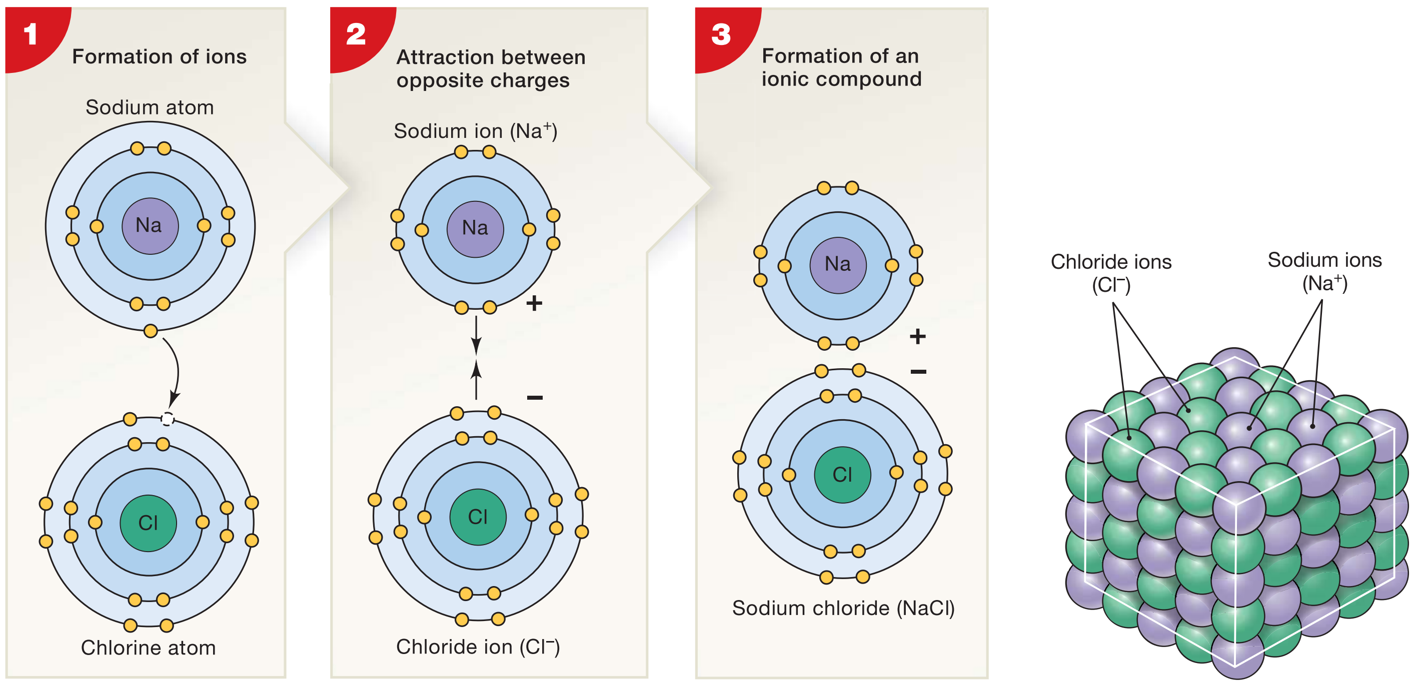
Both atoms have now become stable ions with filled outermost energy levels. But the two ions do not move apart after the electron transfer, because the positively charged sodium ion is attracted to the negatively charged chloride ion (2). The combination of oppositely charged ions forms an *ionic compound*—in this case, **sodium chloride,** otherwise known as table salt (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 2–3b**). Note that ionic compounds, because they consist of an aggregation of ions rather than covalently bonded atoms, are not called molecules. Although sodium chloride and other ionic compounds are common in body fluids, 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 outer electron shells not by gaining or losing electrons, but by sharing electrons with other atoms. Such sharing creates **covalent** (kō-VĀL-ent) **bonds** between the atoms involved.

Individual hydrogen atoms, as diagrammed in **Figure 2–2a**, do not exist in nature. Instead, we find hydrogen molecules.

Figure 2–3 The Formation of Ionic Bonds.



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

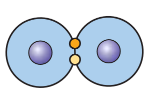
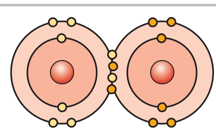
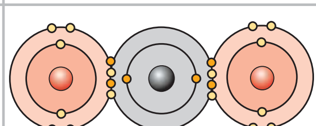
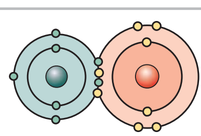
b **Sodium chloride crystal.** Large numbers of sodium and chloride ions form a crystal of sodium chloride (table salt).

Molecular hydrogen consists of a pair of hydrogen atoms (**Figure 2–4**). In chemical shorthand, molecular hydrogen is indicated by 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. The two hydrogen atoms share their electrons, and each electron whirls around both nuclei. The sharing of one pair of electrons creates a **single covalent bond** (—).

Oxygen, with an atomic number of 8, has two electrons in its first energy level and 6 in its second. The oxygen atoms diagrammed in **Figure 2–4** attain a stable electron configuration by sharing 2 pairs of electrons, thereby forming a **double covalent bond.** In a structural formula, a double covalent bond is represented by two lines (═). Molecular oxygen (O_2) is an atmospheric gas that is very important to most organisms. 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.

Figure 2–4 **Covalent Bonds in Four Common Molecules.** In a hydrogen molecule, two hydrogen atoms share electrons such that each atom has a filled outermost electron shell. This sharing creates a single covalent bond. In an oxygen molecule, two oxygen atoms share two pairs of electrons. The result is a double covalent bond. In a carbon dioxide molecule, a central carbon atom forms double covalent bonds with two oxygen atoms. A nitric oxide molecule is held together by a double covalent bond, but the outer electron shell of the nitrogen atom requires an additional electron to be complete. Thus, nitric oxide is a free radical, which reacts readily with another atom or molecule.

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

A triple covalent bond, such as the one joining two nitrogen molecules (N_2), is indicated by three lines (≡). 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 will discuss the reasons for eliminating nitrogen under these conditions in Chapter 23.) Covalent bonds usually form molecules in which the outer energy levels of the atoms involved are complete. An ion or molecule that contains unpaired electrons in its outermost energy level 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. Free radicals sometimes form in the course of normal metabolism, but cells have several methods of removing or inactivating them. However, *nitric oxide* (NO) is a free radical that has important functions in the body. 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. Evidence suggests that the cumulative damage produced by free radicals inside and outside our cells is a major factor in the aging process.

Tips & Tricks

Remember this mnemonic for the 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 one vertical bond), and ***c***arbon shares *4* pairs (–C– with two vertical bonds).

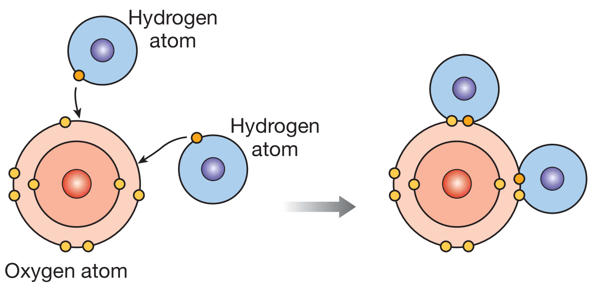
Nonpolar Covalent Bonds

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—which occur, for instance, between two atoms of the same type—are called **nonpolar covalent bonds.** Nonpolar covalent bonds are very common; those involving carbon atoms form most of the structural components of the human body.

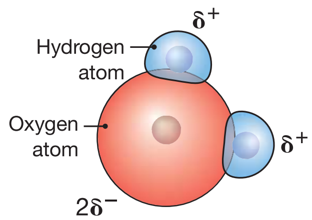
Polar Covalent Bonds

Covalent bonds involving different types of atoms may instead 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.** For example, in a molecule of water (**Figure 2–5**), an oxygen atom forms covalent bonds with two hydrogen atoms. The oxygen nucleus has a much stronger at-

Figure 2–5 Polar Covalent Bonds and the Structure of Water.



a Formation of a water molecule. In forming a water molecule, an oxygen atom completes its outermost energy level 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 Charges on a water molecule. Because the oxygen atom has two 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.

traction for the shared electrons than the hydrogen atoms do. 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^-. 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. Polar molecules have very interesting properties; we will consider the characteristics of the most important polar molecule in the body, water, in a later section.

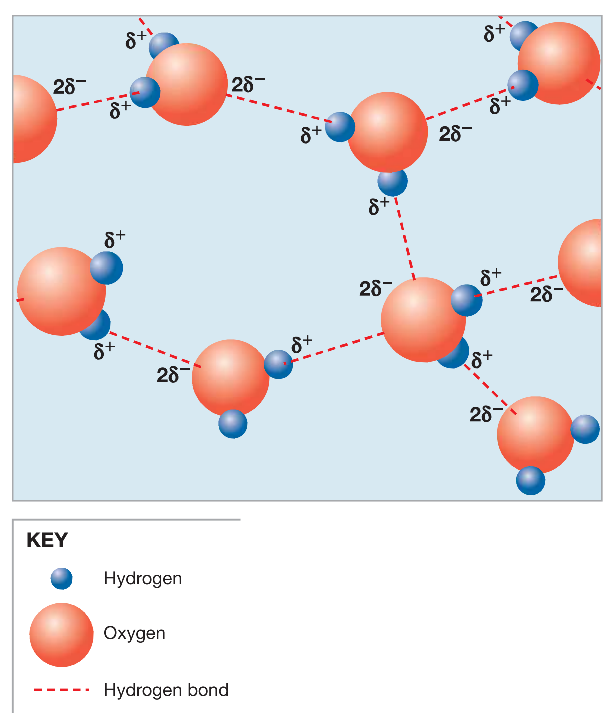
Hydrogen Bonds

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 δ^+ on the hydrogen atom of a polar covalent bond and a δ^- on an oxygen, nitrogen, or fluorine atom of another polar covalent bond. The polar covalent bond containing the oxygen, nitrogen, or fluorine atom can be in a different molecule from, or in the same molecule as, the hydrogen atom. Hydrogen bonds are too weak to create molecules, but they can change molecular shapes or pull molecules together. For example, hydrogen bonding occurs between water molecules (**Figure 2–6**). At a water surface, this attraction between molecules slows the rate of evaporation and creates the phenomenon known as surface tension. **Surface tension** acts as a barrier that keeps small objects from entering the water. For example, it allows insects to walk across the surface of a pond or puddle. Similarly, small objects such as dust particles are prevented from touching the surface of the eye by the surface tension in a layer of tears. At the cellular level, hydrogen bonds affect the shapes and properties of complex molecules, such as proteins and nucleic acids (including DNA), and they may also determine the three-dimensional relationships between molecules.

States of Matter

Most matter in our environment exists in one of three states: solid, liquid, or gas. *Solids* maintain their volume and their shape at ordinary temperatures and pressures. A lump of

Figure 2–6 Hydrogen Bonds 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 2–5b*). The distances between these molecules have been exaggerated for clarity.



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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 neither a constant volume nor a fixed shape. Gases can be compressed or expanded; unlike liquids they will fill a container of any size. The air of our atmosphere is the gas with which we are most familiar.

Whether a particular substance is a solid, a liquid, or a gas depends on the degree of interaction among its atoms or molecules. The particles of a solid are placed 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 in the liquid state over a broad range of temperatures primarily because of hydrogen bonding among the water molecules. We will talk more about the unusual properties of water in a later section.

Molecular Weights

The **molecular weight** of a molecule is the sum of the atomic weights of its component atoms. It follows from the definition of the mole given previously that the molecular weight of a molecule in grams is equal to the weight of one mole of molecules. Molecular weights are important because you can neither handle individual molecules nor easily count the billions of molecules involved in chemical reactions in the body. Using molecular weights, you can calculate the quantities of reactants needed to perform a specific reaction and determine the amount of product generated. For example, suppose you want to create water from hydrogen and oxygen according to the equation

$$2H_2 + O_2 \rightarrow 2H_2O$$

The first step would be to calculate the molecular weights involved. The atomic weight of hydrogen is close to 1.0, so one hydrogen molecule (H_2) has a molecular weight near 2.0. Oxygen has an atomic weight of about 16, so the molecular weight of an oxygen molecule (O_2) is about 32. Thus you would combine 4 g of hydrogen with 32 g of oxygen to produce 36 g of water. You could also work with ounces, pounds, or tons, as long as the proportions remained the same.

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 chemical reactions in physiology

Cells remain 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 occur as atoms in the reacting substances, or **reactants,** are rearranged to form different substances, or **products** (**Spotlight Figure 2–7**).

In effect, each cell is a chemical factory. Growth, maintenance and repair, secretion, and contraction all involve complex chemical reactions. Cells use chemical reactions to provide the energy needed to maintain homeostasis and to perform essential functions. All of the reactions under way in the cells and tissues of the body at any given moment constitute its **metabolism** (me-TAB-ō-lizm).

Basic Energy Concepts

An understanding of some basic relationships between matter and energy is essential 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 like 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 perform work; movement or physical change cannot occur unless energy is provided. The two major types of energy are kinetic energy and potential energy:

1. **Kinetic energy** is the energy of motion—energy that can be transferred to another object and perform 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;

Spotlight

Figure 2–7

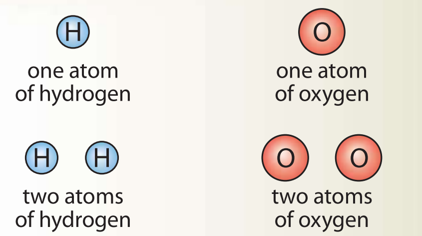
Chemical Notation

Before we can consider the specific compounds that occur in the human body, we must be able to describe chemical compounds and reactions effectively. The use of sentences to describe chemical structures and events often leads to confusion. A simple form of "chemical shorthand" makes communication much more efficient. The chemical shorthand we will use is known as chemical notation. Chemical notation enables us to describe complex events briefly and precisely; its rules are summarized below.

VISUAL REPRESENTATION = **CHEMICAL NOTATION**

Atoms

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



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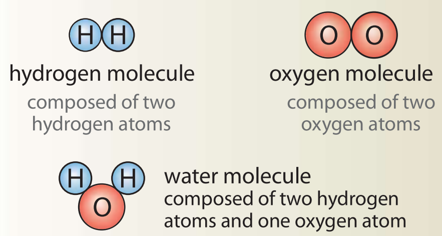
- H — one atom of hydrogen
- O — one atom of oxygen
- $2H$ — two atoms of hydrogen
- $2O$ — two atoms of oxygen

Molecules

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

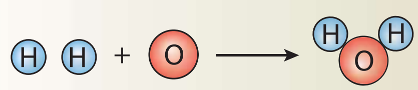


=

- H_2 — hydrogen molecule
- O_2 — oxygen molecule
- H_2O — water molecule

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. An arrow indicates the direction of the reaction, from reactants (usually on the left) to products (usually on the right). In the following reaction, two atoms of hydrogen combine with one 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.

=

$2H + O \longrightarrow H_2O$

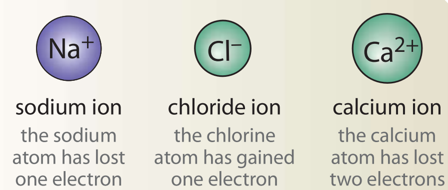
Balanced equation

$2H + 2O \longrightarrow H_2O$

Unbalanced equation

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 lost one electron.) A single minus sign indicates an anion with a charge of –1. (The original atom has gained one electron.) If more than one electron has been lost or gained, the charge on the ion is indicated by a number preceding the plus or minus sign.



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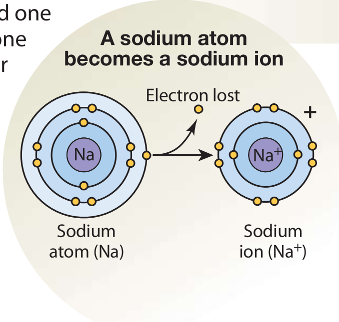
- Na^+ — sodium ion
- Cl^- — chloride ion
- Ca^{2+} — calcium ion



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 perform work.

Cells perform work as they 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.

Decomposition Reactions

Decomposition is a reaction that breaks a molecule into smaller fragments. You could diagram a simple *decomposition reaction* as:

$$AB \longrightarrow A + B$$

Decomposition reactions occur 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 before absorption begins.

Decomposition reactions involving water are important in the breakdown of complex molecules in the body. In **hydrolysis** (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:

$$A - B + H_2O \rightarrow A\text{-}H + HO - B$$

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 perform work. By harnessing the energy released in this way, cells perform 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 could be diagrammed as:

$$A + B \longrightarrow AB$$

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.

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

$$A - H + HO - B \longrightarrow A - B + H_2O$$

Dehydration synthesis is therefore the opposite of hydrolysis. We will encounter examples of both reactions in later sections.

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). Because it takes energy to create a chemical bond, anabolism is usually considered an "uphill" process. Cells must balance their energy budgets, with catabolism providing the energy to support anabolism and other vital functions.

Tips & Tricks

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:

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

Although the reactants and products contain the same components (A, B, C, and D), 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

Chemical reactions are (at least theoretically) 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 occurring simultaneously, one a synthesis ($A + B \longrightarrow AB$) and the other a decomposition ($AB \longrightarrow A + B$). At equilibrium, the rates at which the two reactions proceed are in balance. As fast as one molecule of AB forms, another degrades into A + B.

The result of a disturbance in the equilibrium condition can be predicted. In our example, the rate at which the synthesis reaction proceeds 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. The addition of 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, an equilibrium is again established.

Tips & Tricks

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

Checkpoint

7. **The chemical shorthand used to describe chemical compounds and reactions effectively is known as ___________.**
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. How would you classify this reaction?**

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

2-4 Enzymes catalyze specific biochemical reactions by lowering a reaction's activation energy

Most chemical reactions do not occur spontaneously, or they occur so slowly that they would be of little value to 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.** Although many reactions can be activated by changes in temperature or acidity, 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 perform most of the complex synthesis and decomposition reactions in your body.

Enzymes promote chemical reactions by lowering the activation energy requirements (**Figure 2–8**). In doing so, they make it possible for chemical reactions, such as the breakdown of sugars, to proceed under conditions compatible with life. Enzymes belong to a class of substances called **catalysts** (KAT-uh-lists; *katalysis*, dissolution), compounds that accelerate chemical reactions without themselves being permanently changed or consumed. A cell makes an enzyme molecule to promote a specific reaction. Enzymatic reactions, which are reversible, can be written as

$$A + B \xrightleftharpoons{\text{enzyme}} AB$$

Although the presence of an appropriate enzyme can accelerate a reaction, 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 an equilibrium is reached.

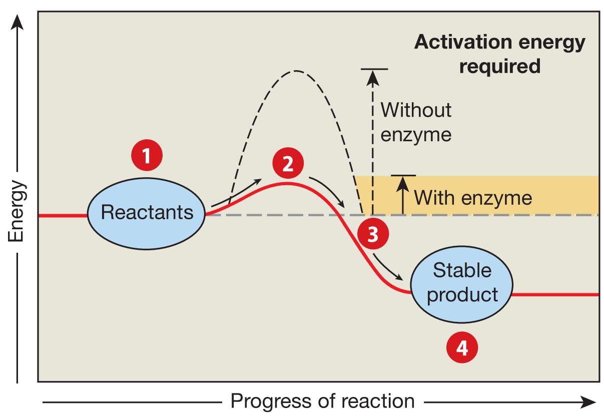
The complex reactions that support life proceed in a series of interlocking steps, each controlled by a specific enzyme. Such a reaction sequence is called a *metabolic pathway*. A synthetic pathway can be diagrammed as

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

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

Figure 2–8 The Effect of Enzymes on Activation Energy. Enzymes lower the activation energy required for a reaction to proceed readily (in order, from 1–4) under conditions in the body.



reaction, there will be a net release of energy. Reactions that release energy are said to be **exergonic** (*exo-*, outside + *ergon*, work). 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). Exergonic reactions are relatively common in the body; they generate the heat that maintains your body temperature.

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

Although the human body is very complex, it contains relatively few elements (Table 2–1, p. 28). But knowing the identity and quantity of each element in the body will not help you understand the body any more than studying the alphabet will help you understand this textbook. Just as 26 letters can be combined to form thousands of different words in this book, 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 constitute the framework of the body and carry on all its life processes. So it is impossible to understand the structure and functioning of the human body without learning about the major classes of chemical compounds.

We will next turn our attention to nutrients and metabolites. **Nutrients** are the essential elements and molecules normally obtained from the diet. **Metabolites** (me-TAB-ō-līts; *metabole*, change), a much larger group, include all the molecules (nutrients included) 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. **Inorganic compounds** generally do not contain carbon and hydrogen atoms as their primary structural ingredients, whereas carbon and hydrogen always form the basis for **organic compounds.**

The most important inorganic compounds in the body are (1) carbon dioxide, a by-product of cell metabolism; (2) oxygen, an atmospheric gas required in important metabolic reactions; (3) water, which accounts for most of our body weight; and (4) inorganic acids, bases, and salts—compounds held together partially or completely by ionic bonds. In this section, we will 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, and all the inorganic acids, bases, and salts we will discuss are dissolved in body fluids.

Checkpoint

13. Compare organic compounds to inorganic compounds.

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

2-6 Physiological systems depend on water

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

Although water is familiar to everyone, it has some highly unusual properties. These properties are a direct result of the hydrogen bonding that occurs between nearby water molecules.

1. *Solubility.* A remarkable number of inorganic and organic molecules are 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 medium in which other atoms, ions, or molecules are dispersed is called the **solvent;** the dispersed substances are the **solutes.** In *aqueous solutions*, water is the solvent. The solvent properties of water are so important that we will consider them further in the next section.
2. *Reactivity.* In our bodies, chemical reactions occur in water, and water molecules are also participants in some reactions. Hydrolysis and dehydration synthesis are two examples noted earlier in the chapter.
3. *High Heat Capacity.* **Heat capacity** is the ability to absorb and retain heat. 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 before all the hydrogen bonds are broken between individual water molecules and they have enough energy to break free and become water vapor, a gas. Consequently, water stays in the liquid state 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 finally does change from a liquid to a gas. This feature accounts for 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*. Because water accounts for up to two-thirds of the weight of the human body, thermal inertia helps stabilize body temperature.
4. *Lubrication.* Water is an effective lubricant because there is little friction between water molecules. So, if even a thin layer

of water separates two opposing surfaces, friction between those surfaces will be greatly reduced. (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 ventral body cavities prevents friction between internal organs, such as the heart or lungs, and the body wall. ⮌ p. 20

The Properties of Aqueous Solutions

Water's chemical structure makes it an unusually effective solvent (**Figure 2–9**). The 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 poles (**Figure 2–9a**). A water molecule is therefore called a **polar molecule,** or a *dipole*.

Many inorganic compounds are held together partly or completely by ionic bonds. In water, these compounds undergo **dissociation** (di-sō-sē-Ā-shun), or **ionization** (ī-on-i-ZĀ-shun). In this process, ionic bonds are broken as the individual ions interact with the positive or negative ends of polar water molecules (**Figure 2–9b**). The result is a mixture of cations and anions surrounded by water molecules. The water molecules around each ion form a *hydration sphere*.

An aqueous solution containing anions and cations will conduct an electrical current. Cations (+) move toward the negative side, or negative terminal, and anions (−) move toward the positive terminal. 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. Chapters 10 and 12 will discuss these processes in more detail.

Electrolytes and Body Fluids

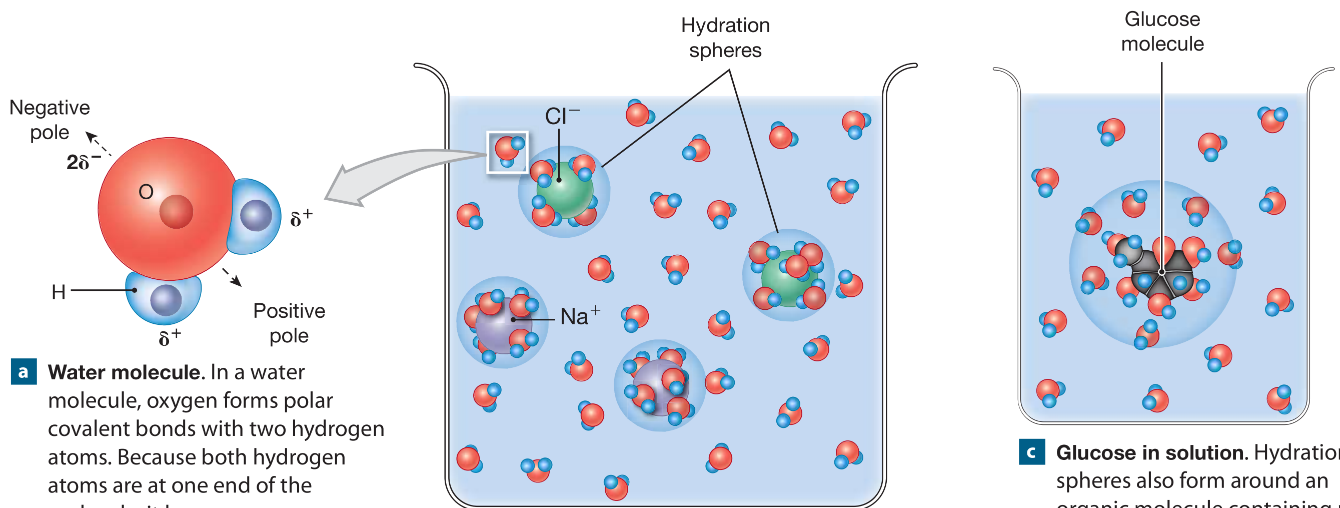
Soluble inorganic molecules whose ions will conduct an electrical current in solution are called **electrolytes** (e-LEK-trō-līts). Sodium chloride is an electrolyte. The dissociation of electrolytes in blood and other body fluids releases a variety of ions. **Table 2–2** lists important electrolytes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

Changes in the concentrations of electrolytes in body fluids will disturb almost every vital function. For example, declining potassium levels will lead to a general muscular paralysis, and rising concentrations will cause weak and irregular heartbeats. The concentrations of ions in body fluids are 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 (**Figure 2–9c**). Molecules such as glucose, an important soluble sugar, that interact readily with water molecules in this way are called **hydrophilic** (hī-drō-FIL-ik; *hydro-*, water + *philos*, loving).

Figure 2–9 The Activities of Water Molecules in Aqueous Solutions.



a Water molecule. In a water molecule, oxygen forms polar covalent bonds with two hydrogen atoms. Because both hydrogen atoms are at one end of the molecule, it has an uneven distribution of 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.

Clinical Note — Solute Concentrations

What's in a mole?

The **concentration** of a substance is the amount of that substance in a specified volume of solvent. Physiologists and clinicians often monitor inorganic and organic solute concentrations in body fluids such as blood or urine. Each solute has a normal range of values (see Appendix), and variations outside this range may indicate disease. Many solutes are reactants or products in biochemical reactions, and as noted earlier, their concentrations directly affect reaction rates.

Solute concentrations can be expressed in several ways. In one method, we express the number of solute atoms, molecules, or ions in a specific volume of solution. Values are reported in moles per liter (mol/L, or M) or millimoles per liter (mmol/L, or mM). A concentration expressed in these units is referred to as the *molarity* of the solution. (Recall that a mole is a quantity of any substance having a weight in grams equal to the atomic or molecular weight of that substance.) Physiological concentrations are most often reported in millimoles per liter.

You can report concentrations in terms of molarity only when you know the molecular weight of the ion or molecule in question. When the chemical structure is unknown or when you are dealing with a complex mixture of materials, concentration is expressed in terms of the weight of material dissolved in a unit volume of solution. Values are then reported in milligrams (mg) or grams (g) per deciliter (dL, or 100 mL). This is the method used, for example, in reporting the concentration of plasma proteins in a blood sample.

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)	$\rightarrow 2Na^+ + SO_4^{2-}$

Many other organic molecules either lack polar covalent bonds or have very few. Such molecules do not have positive and negative terminals, and are said to be *nonpolar*. 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). Among the most familiar hydrophobic molecules are fats and oils of all kinds. 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 discharged into lakes or oceans, they form tenacious oil slicks instead of dissolving.

Tips & Tricks

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 and protein complexes, that are held in solution by their association with water molecules (**Figure 2–9c**). A solution containing dispersed proteins or other large molecules is called a **colloid.** Liquid Jell-O is a familiar viscous (thick) colloid.

The particles or molecules in a colloid will remain in solution indefinitely. A **suspension** contains large particles in solution; if undisturbed, these particles will settle out of solution due to the force of gravity. 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

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 break chemical bonds, change the shapes of complex molecules, and generally 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 that can be represented as:

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

The dissociation of one water molecule yields a hydrogen ion and a *hydroxide* (hī-DROK-sīd) *ion*, OH^-.

Very few water molecules ionize in pure water, and 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 a special shorthand is used to express it. The **pH** of a solution is defined as the negative logarithm of the hydrogen ion concentration in moles per liter. Thus, 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 ($[H^+]=1\times10^{-6}$, or 0.000001) means that the concentration of hydrogen ions is *10 times as great* as it is at a pH of 7 ($[H^+]=1\times10^{-7}$, or 0.0000001). The pH scale ranges from 0 to 14 (**Figure 2–10**).

Although pure water has a pH of 7, 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 pH above 7 is **basic,** or *alkaline* (AL-kuh-lin), meaning that it has more hydroxide ions than hydrogen ions.

The pH of blood normally 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. *Alkalosis* results from an abnormally high pH (above 7.45); a blood pH above 7.8 generally causes uncontrollable and sustained skeletal muscle contractions.

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 are inorganic compounds with important physiological roles

The body contains both inorganic and organic *acids* and *bases* that may cause acidosis or alkalosis, respectively. An **acid** is any solute that dissociates in solution and releases hydrogen

2

ions, thereby lowering the pH. Because a hydrogen atom that loses its electron consists solely of a proton, hydrogen ions are often referred to simply as protons, and acids as *proton donors*.

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 assist in the breakdown of food. Hardware stores sell HCl under the name muriatic acid, for cleaning concrete and swimming pools.

A **base** is a solute that removes hydrogen ions from a solution and thereby acts as a *proton acceptor*. In solution, many bases release a hydroxide ion (OH^-). Hydroxide ions have a strong affinity 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 (Drano) and lye are two familiar examples.

Weak acids and *weak bases* do not dissociate completely. At equilibrium, a significant number of molecules remains intact in the solution. For the same number of molecules in solution, weak acids and weak bases have less of an 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 (table salt) dissociates immediately in water, releasing Na^+ and Cl^-. Sodium and chloride 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 inorganic compounds. Ionic concentrations in the body are regulated by mechanisms we will describe in Chapters 26 and 27.

The ionization of sodium chloride does not affect the local concentrations of hydrogen ions or hydroxide ions, so NaCl, like many salts, is a "neutral" solute. Through their interactions with water molecules, however, other salts may indirectly affect the concentrations of H^+ and OH^- ions. Thus, the dissociation of some 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 2–9**.

The use of antacids such as Alka-Seltzer provides one example of the type of reaction that occurs in buffer systems. Alka-Seltzer uses 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?**

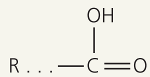
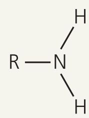
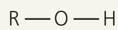
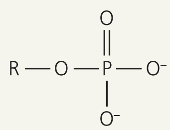
See the blue Answers tab at the back of the book.

2-9 Carbohydrates contain carbon, hydrogen, and oxygen in a 1:2:1 ratio

Carbohydrates are one type of organic compound. Organic compounds always contain the elements carbon and hydrogen, and generally oxygen as well. Many 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.

Many organic molecules are soluble in water. Although the previous discussion focused on inorganic acids and bases, there are also important organic acids and bases. For example, *lactic acid* is an organic acid, generated by active muscle tissues, that must be neutralized by the carbonic acid–bicarbonate buffer system to prevent a potentially dangerous pH decline in body fluids.

Table 2–3 Important Functional Groups of Organic Compounds

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

*The letter R represents the term *R group* and is used to denote the rest of the molecule, whatever that might be. The R group is also known as a side chain.

Although organic compounds are diverse, certain groupings of atoms occur again and again, even in very different types of molecules. These *functional groups* greatly influence the properties of any molecule they are in. **Table 2–3** details the functional groups you will study in this chapter.

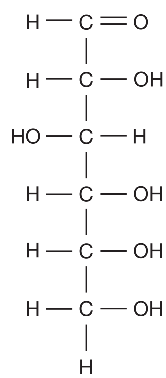
A **carbohydrate** is an organic molecule that contains carbon, hydrogen, and oxygen in a ratio near 1:2:1. Familiar carbohydrates include the *sugars* and *starches* that make up about half of the typical U.S. diet. Carbohydrates typically account for less than 1 percent of total body weight. Although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

Monosaccharides

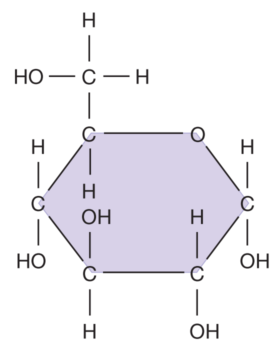
A **monosaccharide** (mon-ō-SAK-uh-rīd; *mono-*, single + *sakcharon*, sugar), or *simple sugar*, is a carbohydrate containing three to seven carbon atoms. A monosaccharide can be called a *triose* (three-carbon), *tetrose* (four-carbon), *pentose* (five-carbon), *hexose* (six-carbon), or *heptose* (seven-carbon). The hexose **glucose** (GLOO-kōs), $C_6H_{12}O_6$, is the most important metabolic "fuel" in the body. The atoms in a glucose molecule may form either a straight chain (**Figure 2–11a**) or a ring (**Figure 2–11b**). In the body, the ring form is more common. A three-dimensional model shows the arrangement of atoms in the ring most clearly (**Figure 2–11c**).

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

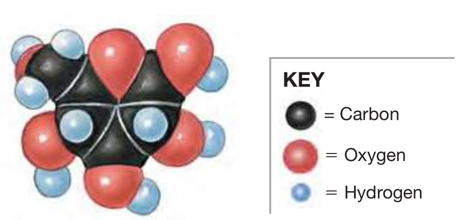
Figure 2–11 The Structure of Glucose.



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 that shows the organization of the atoms in the ring

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. Although its chemical formula, $C_6H_{12}O_6$, is the same as that of glucose, the arrangement of its atoms differs from that of glucose. As a result, separate enzymes and reaction sequences control its breakdown and synthesis. Monosaccharides such as glucose and fructose dissolve readily in water and are rapidly distributed throughout the body by blood and other body fluids.

Disaccharides and Polysaccharides

Carbohydrates other than simple sugars are complex molecules composed of monosaccharide building blocks. Two monosaccharides 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 (**Figure 2–12a**) involves dehydration synthesis, a process introduced earlier in the chapter. Dehydration synthesis, or condensation, links molecules together by the removal of a water molecule. The breakdown of sucrose into simple sugars is an example of hydrolysis, the functional opposite of dehydration synthesis (**Figure 2–12b**).

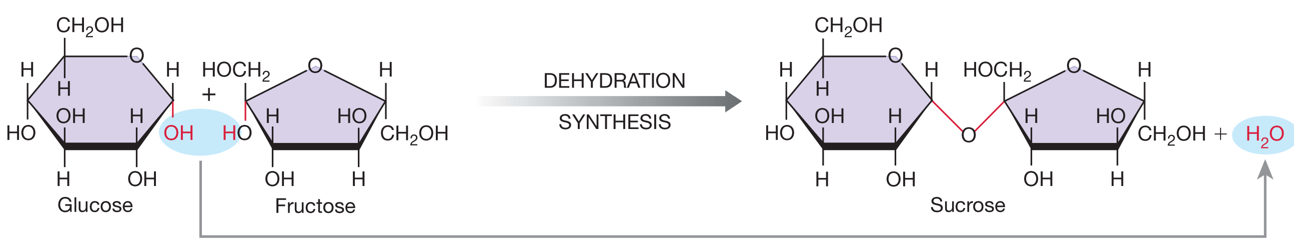
Many foods contain disaccharides, but all carbohydrates except monosaccharides must be disassembled through hydrolysis before they can provide useful energy. Most popular junk foods (high in calories but otherwise lacking in nutritional content), such as candies and sodas, abound in monosaccharides (commonly fructose) and disaccharides (generally sucrose). Some people cannot tolerate sugar for medical reasons; others avoid it in an effort to control their weight. (Excess sugars are converted to fat for long-term storage.) Many such 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 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 bulk to digestive wastes but are useless as a source of energy.

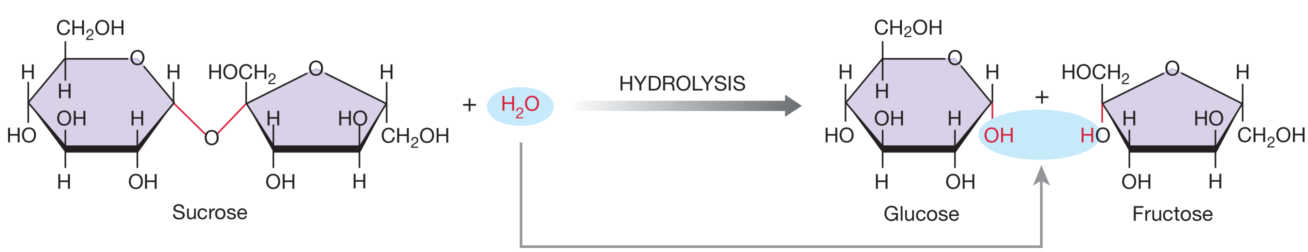
Starches are large polysaccharides formed from glucose molecules. Most starches are manufactured by plants. Your digestive tract can break these molecules into monosaccharides. Starches such as those in potatoes and grains are a major dietary energy source.

The polysaccharide **glycogen** (GLĪ-kō-jen), or *animal starch*, has many side branches consisting of chains of glucose molecules (**Figure 2–13**). 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, they absorb glucose from the blood-stream and rebuild glycogen reserves. **Table 2–4** summarizes information about the carbohydrates.

Figure 2–12 The Formation and Breakdown of Complex Sugars. Enzymes perform both these reactions.



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

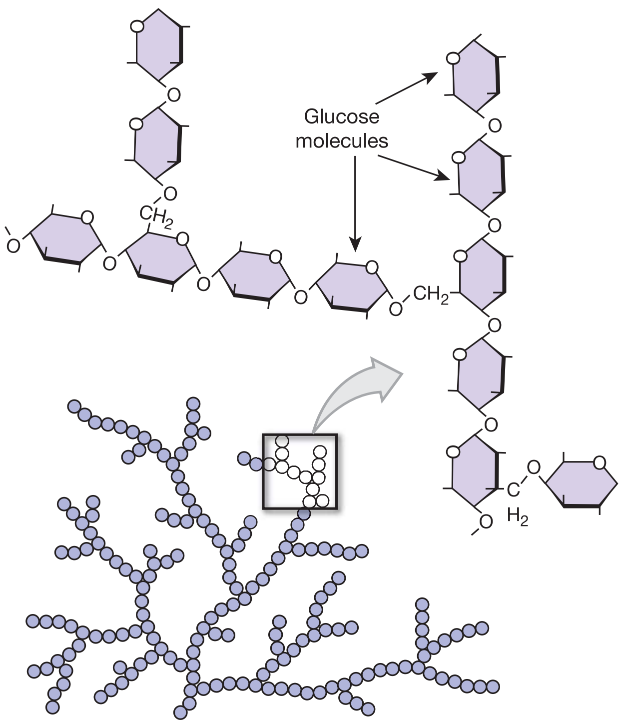


b Breakdown of sucrose into simple sugars by hydrolysis. Hydrolysis reverses the steps of dehydration synthesis; a complex molecule is broken down by the addition of a water molecule.

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

Figure 2–13 The Structure of the Polysaccharide Glycogen. Liver and muscle cells store glucose as the polysaccharide glycogen, a long, branching chain of glucose molecules. This figure uses a different method of representing a carbon ring structure: At five corners of each hexagon is a carbon atom. An oxygen atom occupies the remaining corner in each glucose ring.



Checkpoint

19. **A food contains organic molecules with the elements C, H, and O in a ratio of 1:2:1. What class of compounds do these molecules belong to, and what are their major functions in the body?**

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

2-10 Lipids contain a carbon-to-hydrogen ratio of 1:2

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. The hydrogen-to-oxygen ratio is therefore very large; a representative lipid, such as lauric acid, 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 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 taste and texture as lipids.

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. One end of the carbon chain is always attached to a *carboxyl* (kar-BOK-sil) *group*, —COOH (**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 2–14a** shows a representative fatty acid, *lauric acid*, found in coconut oil and oils of the laurel evergreen.

Clinical Note — Fatty Acids and Health

Good news: The right **fat** can be **good** for you

Humans love fatty foods. Unfortunately, a diet containing large amounts of saturated fatty acids has been shown to increase the risk of heart disease and other cardiovascular problems. Saturated fats are found in such popular foods as fatty meat and dairy products (including such favorites as butter, cheese, and ice cream).

Vegetable oils contain a mixture of monounsaturated and polyunsaturated fatty acids. Recent studies indicate that monounsaturated fats may be more effective than polyunsaturated fats in lowering the risk of heart disease. According to current research, perhaps the healthiest choices are olive and canola oils, which contain particularly abundant quantities of oleic acid, an 18-carbon monounsaturated fatty acid. Surprisingly, compounds called *trans* fatty acids, produced from polyunsaturated oils during the manufacturing of some margarines and vegetable shortenings, appear to increase the risk of heart disease. U.S. Food and Drug Administration (FDA) guidelines now require that *trans* fatty acids be declared in the nutrition label of foods and dietary supplements.

The Inuit people have lower rates of heart disease than do other populations, even though the typical Inuit diet is high in fats and cholesterol. Interestingly, the main fatty acids in the Inuit diet are omega-3s, which means they have an unsaturated bond three carbons before the last (or omega) carbon, a position known as "omega minus 3." Fish flesh and fish oils, a substantial portion of the Inuit diet, contain an abundance of omega-3 fatty acids. Why the presence of omega-3 fatty acids in the diet reduces the risks of heart disease, rheumatoid arthritis, and other inflammatory diseases is not yet apparent; but it is a research topic of great interest.

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, so fatty acids have a very limited solubility in water. In general, the longer the hydrocarbon tail, the lower the solubility of the molecule.

Fatty acids may be either saturated or unsaturated (**Figure 2–14b**). 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 (**Figure 2–14a**). Within the tail, two of those bonds bind adjacent carbon atoms, and the other two bind hydrogen atoms; the carbon atom at the distal 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 has 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 both the shape of the hydrocarbon tail and the way the fatty acid is metabolized. A *monounsaturated* fatty acid has a single double bond in the hydrocarbon tail. A *polyunsaturated* fatty acid contains multiple 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 it cannot be synthesized by the body. The two major classes of eicosanoids are *leukotrienes* and *prostaglandins*. Leukotrienes are produced mostly by cells involved with coordinating the responses to injury or disease. We will consider leukotrienes in Chapters 18 and 22. We consider only prostaglandins here, because virtually all tissues synthesize and respond to them.

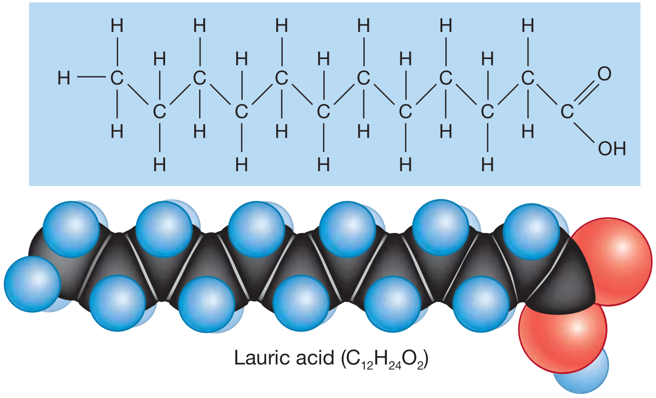
Prostaglandins (pros-tuh-GLAN-dinz) are short-chain fatty acids in which five of the carbon atoms are joined in a ring (**Figure 2–15**). These compounds are released by cells to coordinate or direct local cellular activities, and they are extremely powerful even in small quantities. The effects of prostaglandins 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 but have effects on distant parts are called *hormones*. Hormones are distributed throughout the body in the bloodstream, whereas 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 will discuss hormones and prostaglandins in Chapter 18.

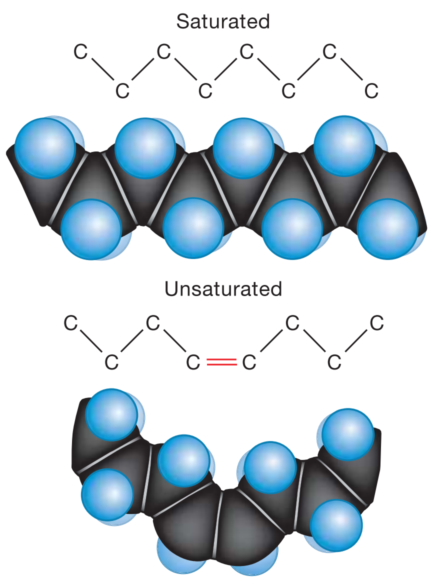
Glycerides

Unlike monosaccharides, individual fatty acids cannot be strung together in a chain by dehydration synthesis. 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 can produce a **monoglyceride** (mon-ō-GLI-ser-īd), consisting of glycerol plus one fatty acid. Subsequent reactions can yield a **diglyceride** (glycerol + two fatty acids) and then a **triglyceride**

Figure 2–14 **Fatty Acids.**

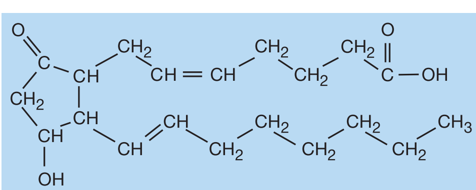


a Lauric acid demonstrates two structural characteristics common to all fatty acids: a long chain of carbon atoms and a carboxyl group (—COOH) at one end.



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.

Figure 2–15 **Prostaglandins.** Prostaglandins are unusual short-chain fatty acids.

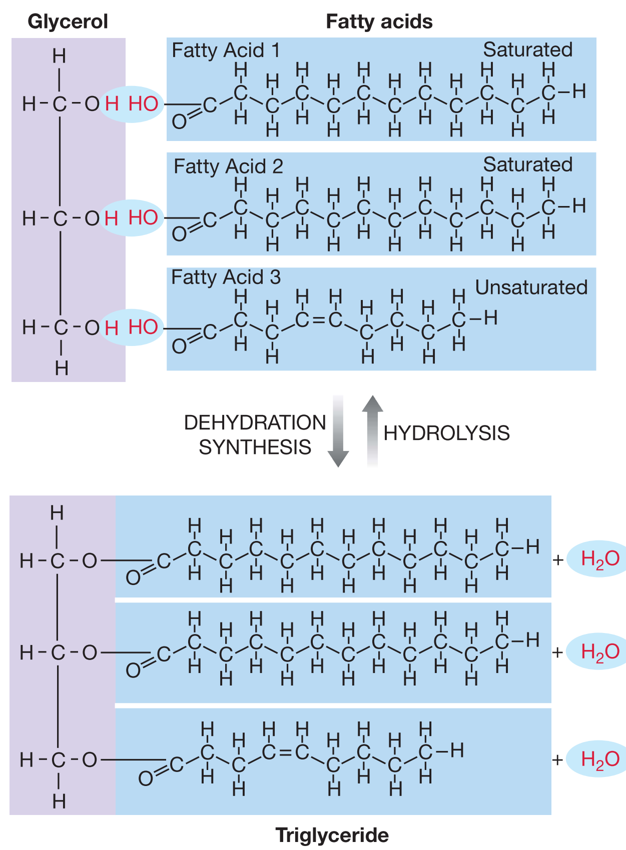


(glycerol + three fatty acids), as in **Figure 2–16**. Hydrolysis breaks the glycerides into fatty acids and glycerol. Comparing **Figure 2–16** with **Figure 2–12** shows that dehydration synthesis and hydrolysis operate the same way, whether the molecules involved are carbohydrates or lipids. Triglycerides, also known as *triacylglycerols* or *neutral fats*, have three important functions.

1. *Energy Source.* Fatty deposits in specialized sites of the body represent a significant energy reserve. In times of need, the triglycerides are disassembled by hydrolysis, yielding fatty acids that can be broken down to provide energy.
2. *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 that through other tissues.
3. *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

Figure 2–16 **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.



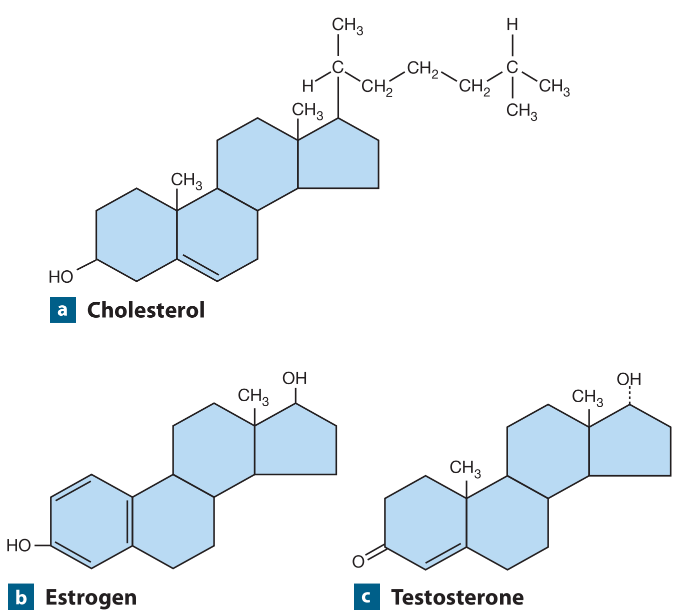
lipid reserves retain both valuable lipid-soluble vitamins (A, D, E, K) and potentially dangerous lipid-soluble pesticides, such as the now-banned DDT.

Steroids

Steroids are large lipid molecules that share a distinctive carbon framework (**Figure 2–17**). They differ in the functional groups that are attached to the basic structure. The steroid **cholesterol** (koh-LES-ter-ol; *chole-*, bile + *stereos*, solid) and related steroids are important for the following reasons:

- The outer boundary of all animal cells, called a plasma membrane, contains cholesterol (**Figure 2–17a**). 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 2–17b,c**).
- Steroid hormones are important in the regulation of tissue metabolism and mineral balance. Examples include the hormones of the adrenal cortex, called *corticosteroids*, and *calcitriol*, 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. Bile salts are produced in the liver and secreted in bile. They interact with lipids in the intestinal tract and assist the digestion and absorption of lipids.

Figure 2–17 **Steroids.** All steroids share a complex four-ring structure. Individual steroids differ in the side chains attached to the carbon rings.



Cholesterol is obtained in two ways: (1) by absorption from animal products in the diet and (2) by synthesis within the body. Liver, fatty meat, cream, and egg yolks are especially rich dietary sources of cholesterol. A diet high in cholesterol can be harmful, because a strong link exists between high blood cholesterol levels and heart disease. Current nutritional advice suggests limiting cholesterol intake to less than 300 mg per day. This amount represents a 40 percent reduction for the average adult in the United States. Unfortunately, because the body can synthesize cholesterol as well, blood cholesterol levels can be difficult to control by dietary restriction alone.

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 2–18a**). In a *glyco*lipid, a carbohydrate is attached to a diglyceride (**Figure 2–18b**). 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 2–18c**). 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.

Cholesterol, phospholipids, and glycolipids 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 hydrophobic lipids. A plasma membrane primarily composed of phospholipids surrounds each cell and separates the aqueous solution inside the cell from the aqueous solution in the extracellular environment. 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 types of lipids and their characteristics are summarized in **Table 2–5**.

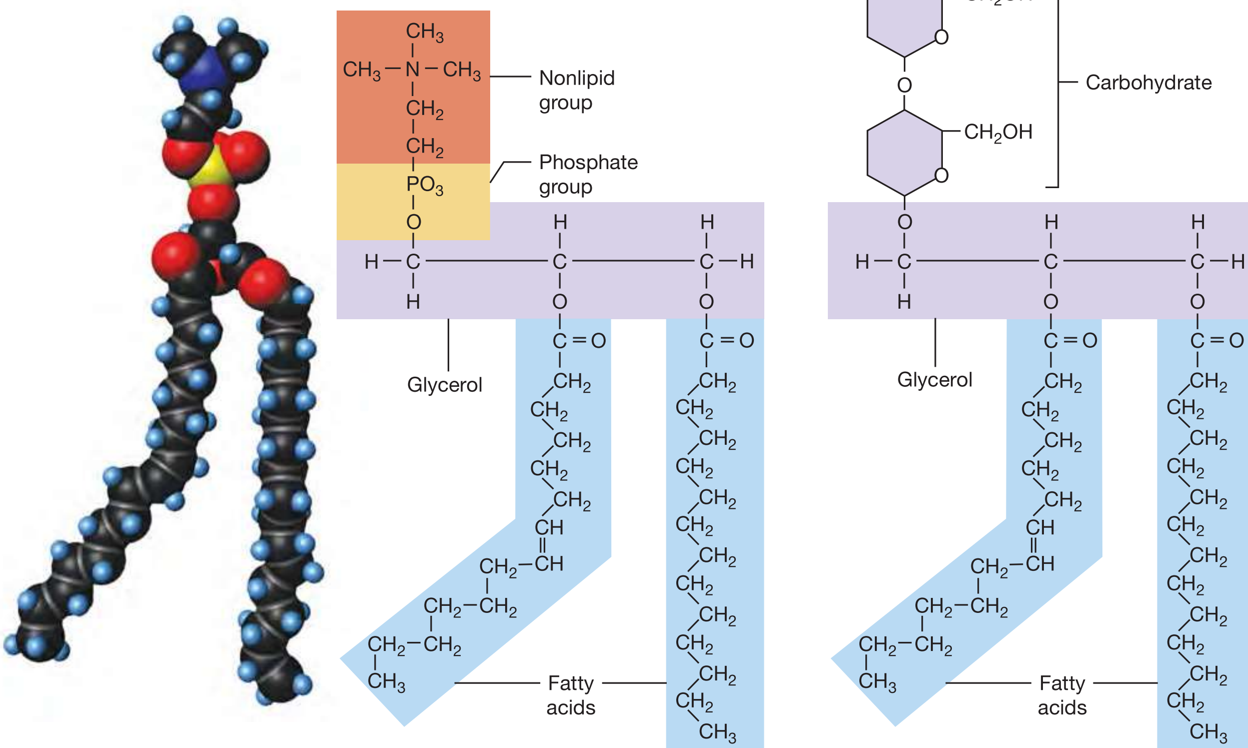
Checkpoint

20. **Describe lipids.**
21. **Which lipids would you find in human plasma membranes?**

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

2

Figure 2–18 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 In large numbers, phospholipids and glycolipids form micelles, with the hydrophilic heads facing the water molecules, and the hydrophobic tails on the inside of each droplet.

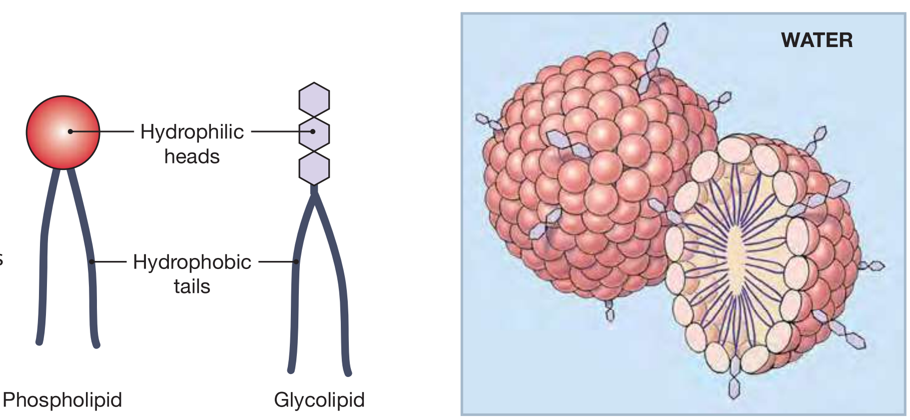


Table 2–5 Representative Lipids and Their Functions in the Body

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 carbon ring framework
Phospholipids, glycolipids	Lecithin (a phospholipid)	Structural components of plasma membranes	Derived from fatty acids and nonlipid components

2-11 Proteins are formed from amino acids and contain carbon, hydrogen, oxygen, and nitrogen

2

Chains of amino acids called **proteins** are the most abundant organic components of 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. Proteins perform a variety of essential functions, which can be classified into seven major categories.

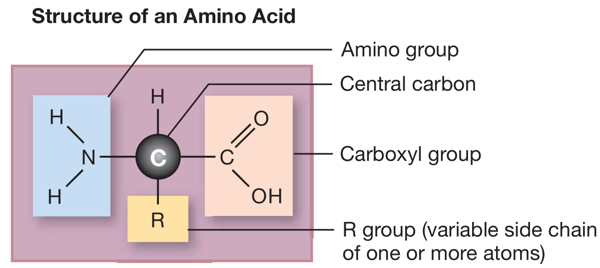
1. *Support. Structural proteins* create a three-dimensional framework for the body, providing strength, organization, and support for cells, tissues, and organs.
2. *Movement. Contractile proteins* are responsible for muscular contraction; related proteins are responsible for the movement of individual cells.
3. *Transport.* Insoluble lipids, respiratory gases, special minerals such as iron, and several hormones cannot be transported in the blood, unless they are first bound to special *transport proteins*. Other specialized proteins transport materials from one part of a cell to another.
4. *Buffering.* Proteins provide a *buffering* action and thereby help prevent dangerous changes in cellular and tissue pH.
5. *Metabolic Regulation. Enzymes* accelerate chemical reactions in cells. The sensitivity of enzymes to environmental factors is extremely important in controlling the pace and direction of metabolic operations.
6. *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.
7. *Defense.* The tough, waterproof proteins of the skin, hair, and nails protect the body from environmental hazards. Proteins called *antibodies*, components of the *immune response*, help protect us from disease. Special *clotting proteins* restrict bleeding after an injury.

Protein Structure

Proteins consist of long chains of organic molecules called **amino acids** (**Figure 2–19**). Twenty different amino acids occur in significant quantities in the body. A typical protein contains 1000 amino acids; the largest protein complexes have 100,000 or more. Each amino acid consists of five components:

- a central carbon atom
- a hydrogen atom
- an *amino group* (—NH_2)
- a *carboxyl group* (—COOH), which can release a hydrogen ion to form a *carboxyl ion* (COO^-)
- an *R group* (a variable *side chain* of one or more atoms)

Figure 2–19 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.



The name *amino acid* refers to the presence of the *amino* group and the acidic carboxyl group, which all amino acids have in common. The different R groups distinguish one amino acid from another, giving each its own chemical properties. However, all 20 amino acids are small, water-soluble molecules.

Protein formation begins as amino acids are strung together to form long chains. **Figure 2–20** 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.

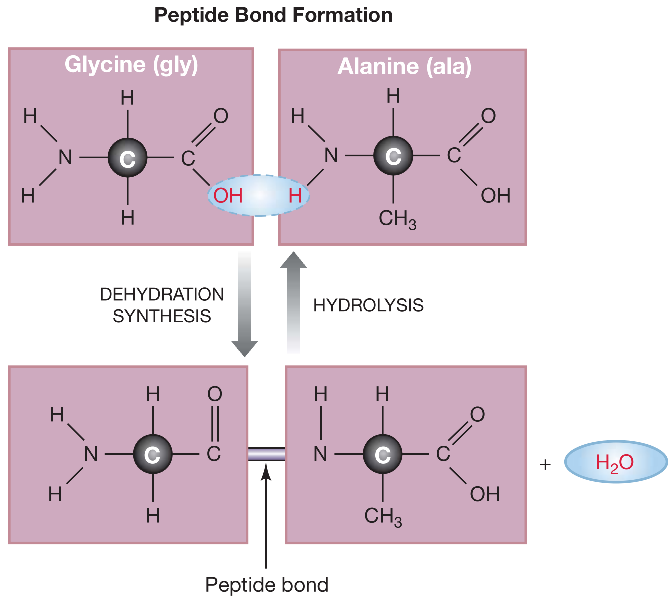
The chain can be lengthened by the addition of more amino acids. Attaching a third amino acid produces a *tripeptide*. Tripeptides and larger peptide chains are called **polypeptides.** Polypeptides containing more than 100 amino acids are usually called proteins. Familiar proteins include *hemoglobin* in red blood cells and *keratin* in fingernails and hair. Because most proteins contain side groups that are negatively charged, the entire protein acts as an anion and is abbreviated Pr^-.

Protein Shape

The characteristics of a particular protein are determined in part by the R groups on its component 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. Proteins can have four levels of structural complexity (**Figure 2–21**).

1. **Primary structure** is the sequence of amino acids along the length of a single polypeptide (**Figure 2–21a**).

Figure 2–20 The Formation of Peptide Bonds. In this example, a peptide bond links the amino acids glycine (for which R = H) and alanine (R = CH_3) to form a dipeptide. Peptides form as dehydration synthesis creates a peptide bond between the carboxyl group of one amino acid and the amino group of another.



2. **Secondary structure** results from bonds between atoms at different parts of the polypeptide chain. Hydrogen bonding, for example, may create either a simple spiral, known as an *alpha-helix*, or a flat *pleated sheet* (**Figure 2–21b**). Which forms depends on the sequence of amino acids in the peptide chain and where hydrogen bonding occurs along the peptide. The alpha-helix is the most common form, but a given polypeptide chain may have both helical and pleated sections.
3. **Tertiary structure** is the complex coiling and folding that gives a protein its final three-dimensional shape (**Figure 2–21c**). Tertiary structure results primarily from interactions between 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 two molecules of the amino acid *cysteine* located at different sites along the chain. 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 2–21d**). Each of the polypeptide subunits has its own secondary and tertiary structures. The protein *hemoglobin* contains four globular subunits. Hemoglobin is found within red blood cells, where it binds and transports oxygen. In *keratin* and *collagen*, three 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. Collagen is the most abundant structural protein and is found in skin, bones, cartilages, and tendons; collagen fibers form the framework that supports cells in most tissues.

Fibrous and Globular Proteins

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

- **Fibrous proteins** form extended sheets or strands. These shapes are usually the product of secondary structure (for proteins that exhibit the pleated-sheet configuration) or quaternary structure (for keratin and collagen). Fibrous proteins are tough, durable, and generally insoluble in water; in the body, they usually play structural roles.
- **Globular proteins** are compact, generally rounded, and readily enter an aqueous solution. The unique shape of each globular protein is the product of its tertiary structure. *Myoglobin*, a protein in muscle cells, is a globular protein, as is hemoglobin, the oxygen-carrying pigment in red blood cells. Many enzymes, hormones, and other molecules that circulate in the bloodstream are globular proteins, as are the enzymes that control chemical reactions inside cells. These proteins can function only if they remain in solution.

Protein Shape and Function

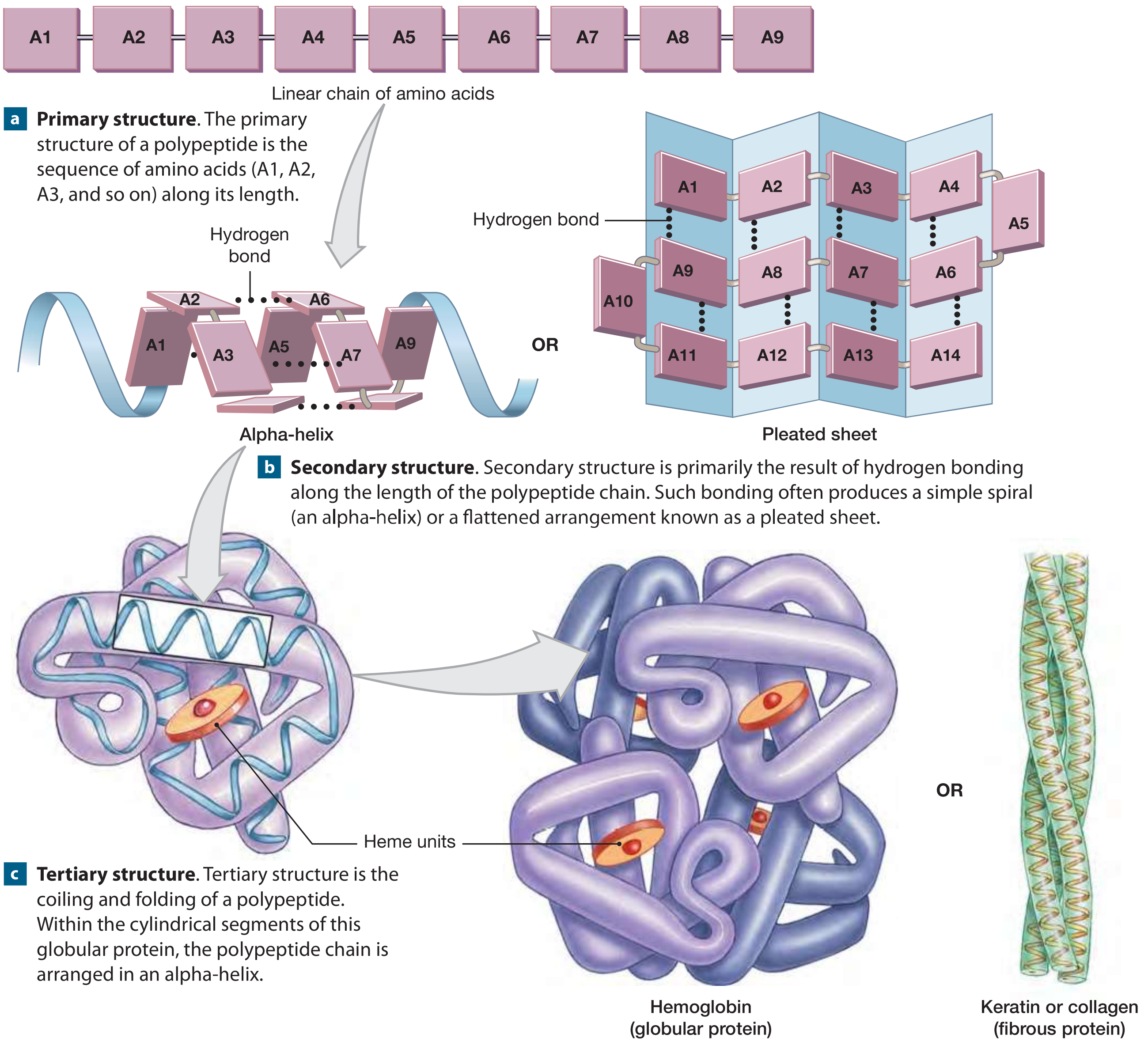
The shape of a protein determines its functional properties, and the ultimate determinant of shape is the sequence of amino acids. The 20 common amino acids can be linked in an astonishing number of combinations, creating proteins of enormously varied shape and function. Changing the identity of only one of the 10,000 or more amino acids in a protein can significantly alter the protein's functional properties. For example, several cancers and *sickle cell anemia*, a blood disorder, result from single changes 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 the function of enzymes, for these proteins are essential to the metabolic operations occurring in every one of our cells.

Enzyme Function

Among the most important of all the body's proteins are the enzymes, first introduced earlier in this chapter. These molecules catalyze the reactions that sustain life: Almost everything that

Figure 2–21 **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 (an alpha-helix) or a flattened arrangement known as a pleated 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.

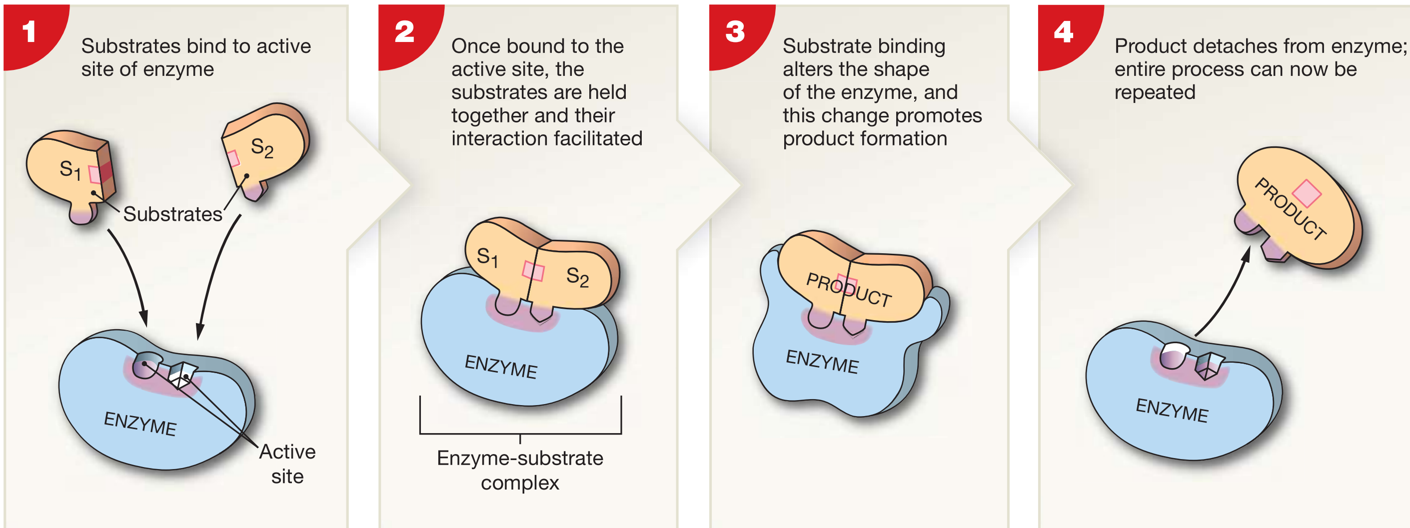
d **Quaternary structure.** Quaternary structure develops when separate polypeptide subunits interact to form a larger molecule. A single hemoglobin molecule contains four globular subunits. Hemoglobin transports oxygen in the blood; the oxygen binds reversibly to the heme units. In keratin and collagen, three fibrous subunits intertwine. Keratin is a tough, water-resistant protein in skin, hair, and nails. Collagen is the principal extracellular protein in most organs.

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—to accelerate a chemical reaction without itself being permanently changed or consumed—the substrates must bind to a special region of the enzyme. This region, called the **active site,** 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 2–22 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).

Figure 2–22 A Simplified View of Enzyme Structure and Function. Each enzyme contains a specific active site somewhere on its exposed surface.



Substrate binding typically results in a temporary, reversible change in the shape of the enzyme that may place physical stresses on the substrate molecules, leading to product formation (3). Product release frees the enzyme, which is then free 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 2–22 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 three basic characteristics:

1. *Specificity.* Each enzyme catalyzes only one type of reaction, a characteristic called **specificity.** An enzyme's specificity is determined by the ability of its active sites to bind only to substrates with particular shapes and charges. Thus, differences in enzyme structure that neither affect the active site nor change the response of the enzyme to substrate binding do not affect enzyme function. In fact, different tissues typically contain enzymes that differ slightly in structure, but catalyze the same reaction. Such enzyme variants are called **isozymes.**
2. *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 encounters will be. 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 have 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.
3. *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 thereby control reaction rates inside the cell. Because the 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 the 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 cofactors include ions such as calcium (Ca^{2+}) and magnesium (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 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.

Effects of Temperature and pH on 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. Death occurs at very high body temperatures (above 43°C, or 110°F) because the denaturation of structural proteins and enzymes soon causes irreparable damage to organs and organ systems. However, this denaturation can be reversed if the temperature is reduced before the individual dies.

Enzymes are equally sensitive to changes in pH. *Pepsin*, an enzyme that breaks down proteins in stomach contents, works best at a pH of 2.0 (strongly acidic). Your small intestine contains *trypsin*, another enzyme that attacks proteins. 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 the identification of normal versus abnormal cells, as well as in the initiation and coordination of 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. The proteoglycans in tissue fluids give them a syrupy consistency.

Checkpoint

22. **Describe a protein.**
23. **How does boiling a protein affect its structural and functional properties?**

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

2-12 DNA and RNA are nucleic acids

Nucleic (noo-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-noo-KLĀ-ik) **acid,** or **DNA,** and **ribonucleic** (rī-bō-noo-KLĀ-ik) **acid,** or **RNA.** As we will see, these two classes of nucleic acids differ in composition, structure, and function.

The DNA in our cells determines our inherited characteristics, including eye color, hair color, and blood type. DNA affects all aspects of body structure and function, because DNA molecules encode the information needed to build proteins. By directing the synthesis of structural proteins, DNA controls the shape and physical characteristics of our bodies. By controlling the manufacture of enzymes, DNA regulates not only protein synthesis, but all aspects of cellular metabolism, including the creation and destruction of lipids, carbohydrates, and other vital molecules.

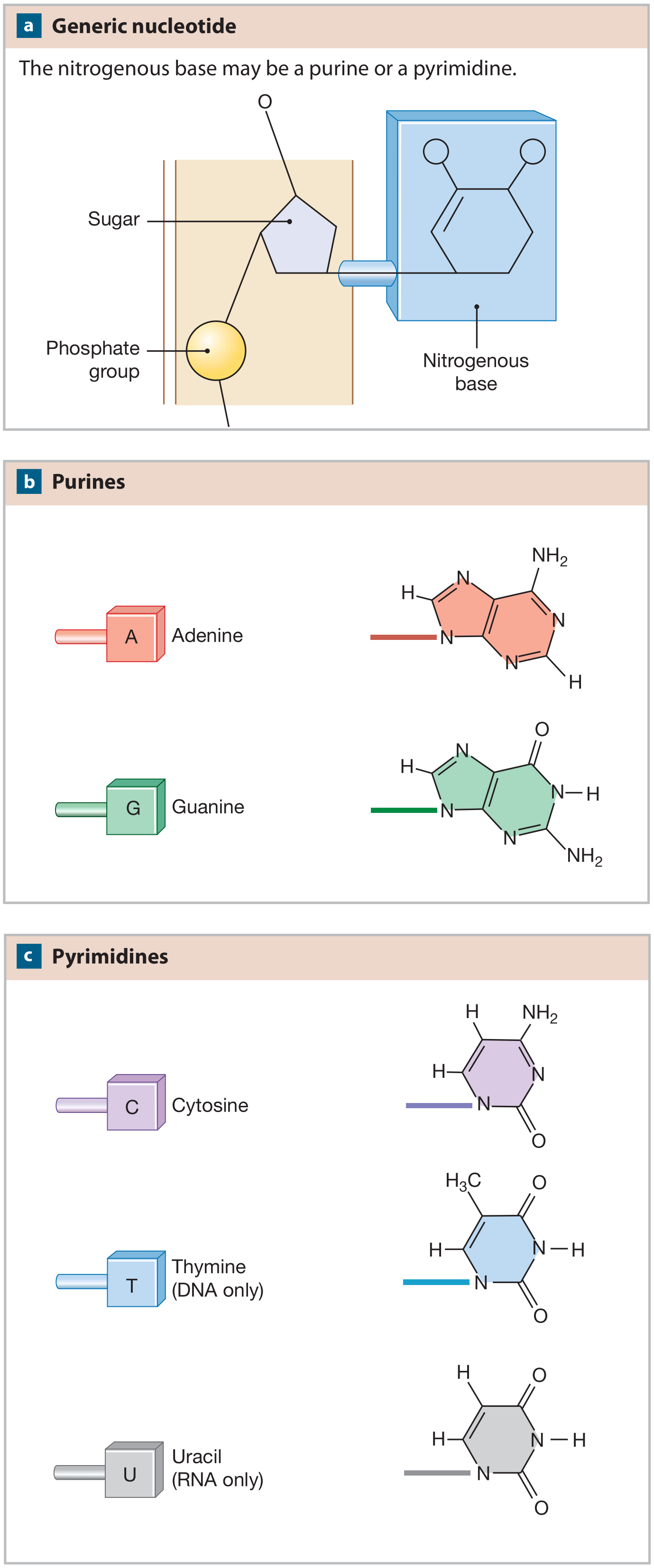
Several forms of RNA cooperate to manufacture specific proteins by using the information provided by DNA. We will detail the functional relationships between DNA and RNA in Chapter 3.

Structure of Nucleic Acids

A nucleic acid consists of one or two long chains that are formed by dehydration synthesis. The individual subunits are called **nucleotides** (**Figure 2–23**). Each nucleotide has three components: (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)** (**Figure 2–23b,c**). Adenine and guanine are double-ringed molecules called *purines*; the other three bases are single-ringed molecules called *pyrimidines*. Both RNA and DNA contain adenine, guanine, and cytosine. Uracil occurs only in RNA and thymine only in DNA.

A nucleotide forms when a phosphate group binds to a pentose already attached to a nitrogenous base. In the formation of a nucleic acid, dehydration synthesis then attaches the phosphate group of one nucleotide to the sugar of another. The "backbone" of a nucleic acid molecule is thus a linear sugar-to-phosphate-to-sugar sequence, with the nitrogenous bases projecting to one side (**Figure 2–24**). The primary role of nucleic acids is the storage and transfer of information—specifically, information essential to the synthesis of proteins within our cells. Regardless of whether we are speaking of DNA or RNA, it is the sequence of nitrogenous bases that carries the information.

Figure 2–23 **Nucleotides and Nitrogenous Bases.**



RNA and DNA

Important structural differences distinguish RNA from DNA. A molecule of RNA consists of a single chain of nucleotides (**Figure 2–24a**). Its shape depends on the order of the nucleotides and the interactions among them. Our cells have three types of RNA: (1) *messenger RNA (mRNA)*, (2) *transfer RNA (tRNA)*, and (3) *ribosomal RNA (rRNA)*. These types have different shapes and functions, but all three are required for the synthesis of proteins, as you will see in Chapter 3.

A DNA molecule consists of a *pair* of nucleotide chains (**Figure 2–24b**). Hydrogen bonding between opposing nitrogenous bases holds the two strands together. The shapes of

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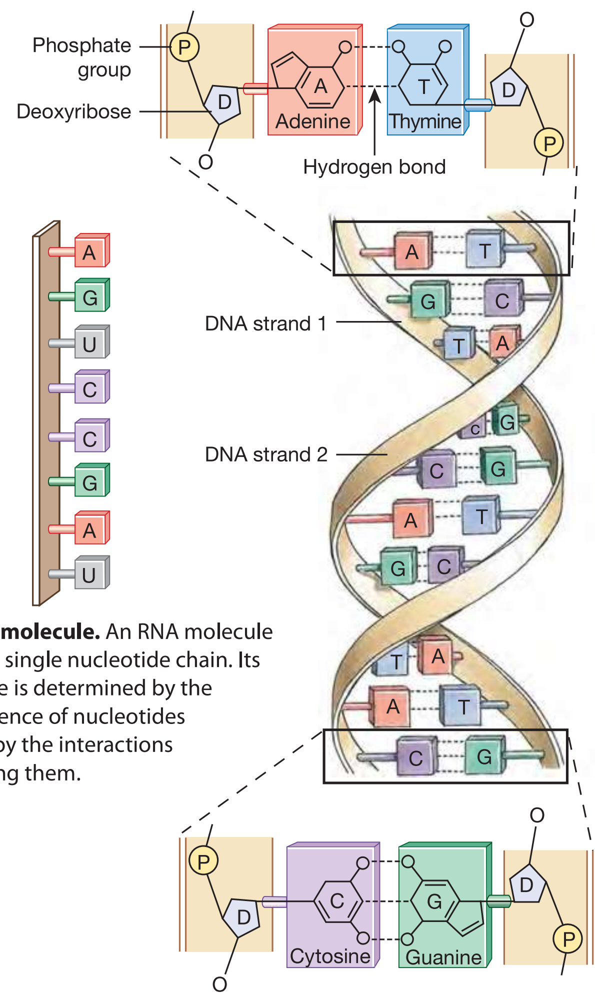
Figure 2–24 **The Structure of Nucleic Acids.** Nucleic acids are long chains of nucleotides. Each molecule starts at the sugar of the first nucleotide and ends at the phosphate group of the last member of the chain.



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.

2

Table 2–6 Comparison of RNA with 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

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.** 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. 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 (**Figure 2–24b**). **Table 2–6** compares RNA with DNA.

Checkpoint

24. **Describe a nucleic acid.**
25. **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.

2-13 ATP is a high-energy compound used by cells

To perform their vital functions, cells must use energy, obtained 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 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 your cells, a high-energy bond generally connects a phosphate group (PO_4^{3-}) to an organic molecule. The resulting product is called a **high-energy compound.** Most high-energy compounds are derived from nucleotides, the building blocks of nucleic acids.

The attachment of a phosphate group to another molecule is called **phosphorylation** (fos-for-i-LĀ-shun). This process does not necessarily produce high-energy bonds. 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*). 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 thereby create the high-energy compound **adenosine triphosphate,** or **ATP** (**Figure 2–25**).

Figure 2–25 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.

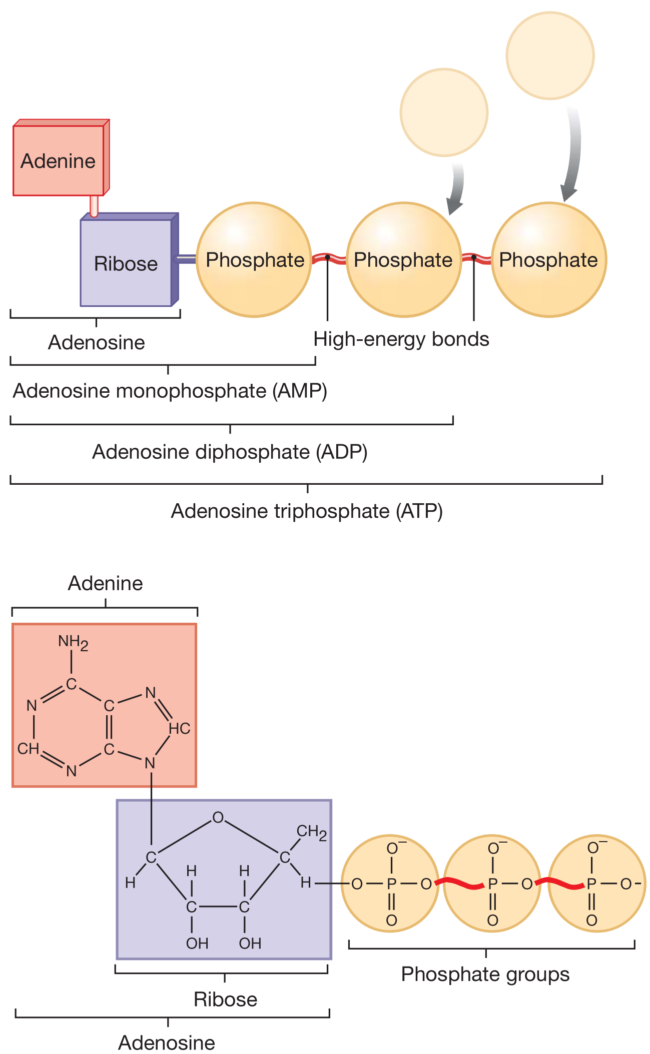


Table 2–7 Classes of Inorganic and Organic Compounds

Class	Building Blocks	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	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	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	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	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

The conversion of ADP to ATP is the most important method of energy storage in our cells; the reversion of ATP to ADP is the most important method of energy release. The relationships involved can be diagrammed as

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

The conversion of ATP to ADP requires an enzyme known as **adenosine triphosphatase,** or **ATPase.** Throughout life, cells continuously generate ATP from ADP and use the energy provided by the ATP to perform vital functions, such as the synthesis of proteins or the contraction of muscles.

Although ATP is the most abundant high-energy compound, there are others—typically, other nucleotides that have undergone phosphorylation. For example, *guanosine triphosphate (GTP)* and *uridine triphosphate (UTP)* are nucleotide-based high-energy compounds that transfer energy in specific enzymatic reactions.

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

Checkpoint

26. **Describe ATP.**
27. **What molecule is produced by the phosphorylation of ADP?**

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

2-14 Chemicals form functional units called cells

The human body is more than a collection of chemicals. The biochemical building blocks discussed in this chapter form functional units called *cells*. p. 6 Each cell behaves like a miniature organism, responding to internal and external stimuli. This is possible only because cells are dynamic structures that adapt to changes in their environment. Such adaptation may involve changes in the chemical organization of the cell—changes that are easily made because organic molecules other than DNA are temporary rather than permanent components of the cell. Their continuous removal and replacement are part of the process of **metabolic turnover.**

Most of the organic molecules in the cell are replaced at intervals ranging from hours to months. The average time between synthesis and breakdown is known as the *turnover time*. **Table 2–8** lists the turnover times of the organic components of representative cells. In the next chapter we will learn more about the functions of these organic components as we explore the cellular level of organization.

2

Table 2–8 Turnover Times		
Cell Type	**Component**	**Average Recycling Time***
Liver	Total protein	5–6 days
	Enzymes	1 hour to several days, depending on the enzyme
	Glycogen	1–2 days
	Cholesterol	5–7 days
Muscle cell	Total protein	30 days
	Glycogen	12–24 hours
Neuron	Phospholipids	200 days
	Cholesterol	100+ days
Fat cell	Triglycerides	15–20 days

*Most values were obtained from studies on mammals other than humans.

Checkpoint

28. **Identify biochemical building blocks discussed in this chapter that are the components of cells.**
29. **Define metabolic turnover.**

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 non-essential and can be harmful to health in excessive amounts. These include cadmium, antimony, chromium, mercury, lead, and arsenic.

hypercholesterolemia: The presence of excess cholesterol in the blood.

radiation sickness: Sickness that results from exposure to radiation and is commonly marked by fatigue, nausea, vomiting, loss of teeth and hair, and in more severe cases by damage to blood-forming tissue.

Chapter Review

Study Outline

An Introduction to the Chemical Level of Organization p. 27

1. Chemicals combine to form complex structures.

2-1 Atoms are the basic particles of matter p. 27

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. *(Figure 2–1)*
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; Table 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 determine an element's chemical properties. *(Figure 2–2)*

2-2 Chemical bonds are forces formed by atom interactions p. 30

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.
8. 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 2–3)*
9. 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.** 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 2–4, 2–5)*
10. A **hydrogen bond** is a weak, but important, force that can affect the shapes and properties of molecules. *(Figure 2–6)*
11. Matter can exist as a *solid*, a *liquid*, or a *gas*, depending on the nature of the interactions among the component atoms or molecules.
12. The molecular weight of a molecule is the sum of the atomic weights of its component atoms.
13. The rules of **chemical notation** are used to describe chemical compounds and reactions. *(Spotlight Figure 2–7)*

2-3 Decomposition, synthesis, and exchange reactions are important chemical reactions in physiology p. 34

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 opposing reactions are in balance.

2-4 ◗ Enzymes catalyze specific biochemical reactions by lowering a reaction's activation energy p. 37

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 requirements. *(Figure 2–8)*
21. **Exergonic** reactions release energy; **endergonic** reactions absorb energy.

2-5 ◗ Inorganic compounds lack carbon, and organic compounds contain carbon p. 38

22. **Nutrients** are the essential elements and molecules normally obtained from the diet; **metabolites** 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. 38

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 2–9)*
25. Many inorganic compounds, called **electrolytes,** undergo **dissociation,** or **ionization,** in water to form ions. *(Figure 2–9; Table 2–2)* Molecules that interact readily with water molecules are called **hydrophilic;** those that do not are called **hydrophobic.**

2-7 ◗ Body fluid pH is vital for homeostasis p. 40

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 2–10)*

2-8 ◗ Acids, bases, and salts are inorganic compounds with important physiological roles p. 41

27. An **acid** releases hydrogen ions; a **base** removes hydrogen ions from a solution. *Strong acids* and *strong bases* ionize completely, whereas *weak acids* and *weak bases* do not.
28. A **salt** is an electrolyte whose cation is not a hydrogen ion (H^+) 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 ◗ Carbohydrates contain carbon, hydrogen, and oxygen in a 1:2:1 ratio p. 42

30. Carbon and hydrogen are the main constituents of **organic compounds,** which generally contain oxygen as well. The properties of the different classes of organic compounds are due to the presence of *functional groups* of atoms. *(Table 2–3)*
31. **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 monosaccharides by **dehydration synthesis.** *(Figures 2–11 to 2–13; Table 2–4)*

2-10 ◗ Lipids contain a carbon-to-hydrogen ratio of 1:2 p. 45

32. **Lipids** include *fats*, *oils*, and *waxes*; most are water-insoluble molecules. The five important classes of lipids are **fatty acids, eicosanoids, glycerides, steroids,** and **phospholipids** and **glycolipids.** *(Figures 2–14 to 2–18; Table 2–5)*
33. **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. *(Figure 2–16)*
34. Steroids (1) are components of plasma membranes, (2) include sex hormones and hormones regulating metabolic activities, and (3) are important in lipid digestion. *(Figure 2–17)*
35. **Phospholipids** and **glycolipids** are structural lipids that are components of *micelles* and plasma membranes.

2-11 ◗ Proteins are formed from amino acids and contain carbon, hydrogen, oxygen, and nitrogen p. 50

36. **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*.
37. Proteins are chains of **amino acids.** Each amino acid 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 2–19, 2–20)*
38. 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). **Fibrous proteins,** such as *keratin* and *collagen*, are elongated, tough, durable, and generally insoluble. **Globular proteins,** such as *myoglobin*, are generally rounded and water-soluble. *(Figure 2–21)*
39. 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 substrate binding can occur. **Coenzymes** are organic cofactors commonly derived from *vitamins*. *(Figure 2–22)*
40. 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** at temperatures or pH values outside the normal range.

2-12 ◗ DNA and RNA are nucleic acids p. 54

41. **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 2–23, 2–24; Table 2–6)*

42. Nucleic acids are 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; it contains **uracil** instead of thymine.

2

2-13 ATP is a high-energy compound used by cells p. 56

43. 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. The cell can use this energy to power essential activities. *(Figure 2–25; Table 2–7)*

2-14 Chemicals form functional units called cells p. 57

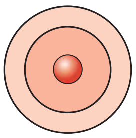
44. Biochemical building blocks form functional units called *cells*.
45. The continuous removal and replacement of cellular organic molecules (other than DNA), a process called **metabolic turnover,** allows cells to change and to adapt to changes in their environment. *(Table 2–8)*

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?

$$A–B–C–D + H_2O \rightarrow A–B–C–H + HO–D$$

3. The lightest of an atom's main components
   (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 determines 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 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. Chemical reactions in the human body are controlled by special catalytic molecules called
    (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 stable fundamental particles in atoms?
17. What four major classes of organic compounds 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) 0
    (b) 5
    (c) 10
    (d) 11
    (e) 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) monoglyceride.
    (d) diglyceride.
    (e) triglyceride.
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?

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