# The Chemical Basis of Life

CHAPTER

2

Colorized scanning electron micrograph (SEM) of bundles of collagen fibers (brown) and elastic fibers (blue). The chemical composition of these fibers determines their functions within the body.

All the structures of the body are composed of chemicals, and all the functions of the body result from the interactions of these chemicals with one another. The generation of nerve impulses and the physiologic processes of digestion, muscle contraction, and metabolism can be described in chemical terms. Likewise, many illnesses and their treatment can be described chemically. For example, Parkinson's disease, which causes uncontrollable shaking movements, results from a shortage of a chemical called dopamine in certain nerve cells in the brain. It is treated by giving patients another chemical that is converted to dopamine by brain cells.

To understand anatomy and physiology, it is essential to have a basic knowledge of chemistry—the scientific discipline concerned with the atomic composition and structure of substances and the reactions they undergo. This chapter outlines *basic chemistry* (27), *chemical reactions and energy* (34), *inorganic chemistry* (39), and *organic chemistry* (43). It is not a comprehensive review of chemistry, but it does review some of the basic concepts. Refer back to this chapter when chemical phenomena are discussed later in the text.

# Basic Chemistry

***Objectives***

- *Define the terms matter, mass, weight, element, and atom.*
- *Describe the subatomic particles of an atom and explain how they determine atomic number, mass number, isotopes, and atomic mass.*
- *Describe the types of chemical bonding and contrast them with intermolecular forces.*
- *Distinguish between a molecule and a compound and describe how each dissolves in water.*

## Matter, Mass, and Weight

All living and nonliving things are composed of **matter,** which is anything that occupies space and has mass. **Mass** is the amount of matter in an object, and **weight** is the gravitational force acting on an object of a given mass. For example, the weight of an apple results from the force of gravity "pulling" on the apple's mass.

**P R E D I C T 1**

**The difference between mass and weight can be illustrated by considering an astronaut. How does an astronaut's mass and weight in outer space compare to the astronaut's mass and weight on the earth's surface?**

The **kilogram (kg),** which is the mass of a platinum–iridium cylinder kept at the International Bureau of Weights and Measurements in France, is the international unit for mass. The mass of all other objects is compared to this cylinder. For example, a 2.2-pound lead weight or 1 liter (L) (1.06 qt) of water each has a mass of approximately 1 kg. An object with 1/1000 the mass of a kilogram is defined as having a mass of 1 **gram (g).**

Chemists use a balance to determine the mass of objects. Although we commonly refer to "weighing" an object on a balance, we are actually "massing" the object because the balance compares objects of unknown mass to objects of known mass. When the unknown and known masses are exactly balanced, the gravitational pull of the earth on both of them is the same. Thus, the effect of gravity on the unknown mass is counteracted by the effect of gravity on the known mass. A balance produces the same results on a mountaintop as at sea level because it does not matter if the gravitational pull is strong or weak. It only matters that the effect of gravity on both the unknown and known masses is the same.

## Elements and Atoms

An **element** is the simplest type of matter with unique chemical properties. To date, 112 elements are known. A list of the elements commonly found in the human body is given in table 2.1. About

**Table 2.1 Common Elements**

| Element | Symbol | Atomic Number | Mass Number | Atomic Mass | Percent in Human Body by Weight (%) | Percent in Human Body by Number of Atoms (%) |
|---|---|---|---|---|---|---|
| Hydrogen | H | 1 | 1 | 1.008 | 9.5 | 63.0 |
| Carbon | C | 6 | 12 | 12.01 | 18.5 | 9.5 |
| Nitrogen | N | 7 | 14 | 14.01 | 3.3 | 1.4 |
| Oxygen | O | 8 | 16 | 16.00 | 65.0 | 25.5 |
| Fluorine | F | 9 | 19 | 19.00 | Trace | Trace |
| Sodium | Na | 11 | 23 | 22.99 | 0.2 | 0.3 |
| Magnesium | Mg | 12 | 24 | 24.31 | 0.1 | 0.1 |
| Phosphorus | P | 15 | 31 | 30.97 | 1.0 | 0.22 |
| Sulfur | S | 16 | 32 | 32.07 | 0.3 | 0.05 |
| Chlorine | Cl | 17 | 35 | 35.45 | 0.2 | 0.03 |
| Potassium | K | 19 | 39 | 39.10 | 0.4 | 0.06 |
| Calcium | Ca | 20 | 40 | 40.08 | 1.5 | 0.31 |
| Chromium | Cr | 24 | 52 | 51.00 | Trace | Trace |
| Manganese | Mn | 25 | 55 | 54.94 | Trace | Trace |
| Iron | Fe | 26 | 56 | 55.85 | Trace | Trace |
| Cobalt | Co | 27 | 59 | 58.93 | Trace | Trace |
| Copper | Cu | 29 | 63 | 63.55 | Trace | Trace |
| Zinc | Zn | 30 | 64 | 65.39 | Trace | Trace |
| Selenium | Se | 34 | 80 | 78.96 | Trace | Trace |
| Molybdenum | Mo | 42 | 98 | 95.94 | Trace | Trace |
| Iodine | I | 53 | 127 | 126.9 | Trace | Trace |

96% of the weight of the body results from the elements oxygen, carbon, hydrogen, and nitrogen.

An **atom** is the smallest particle of an element that has the chemical characteristics of that element. An element is composed of atoms of only one kind. For example, the element carbon is composed of only carbon atoms, and the element oxygen is composed of only oxygen atoms.

An element, or an atom of that element, often is represented by a symbol. Usually the first letter or letters of the element's name are used—for example, C for carbon, H for hydrogen, Ca for calcium, and Cl for chlorine. Occasionally the symbol is taken from the Latin, Greek, or Arabic name for the element—for example, Na from the Latin word *natrium* is the symbol for sodium.

### Atomic Structure

The characteristics of living and nonliving matter result from the structure, organization, and behavior of atoms (figure 2.1). Atoms are composed of subatomic particles, some of which have an electric charge. The three major types of subatomic particles are neutrons, protons, and electrons. **Neutrons** have no electric charge, **protons** have positive charges, and **electrons** have negative charges. The positive charge of a proton is equal in magnitude to the negative charge of an electron. Because equal numbers of protons and electrons occur in an atom, the individual charges cancel each other, and the atom is electrically neutral.

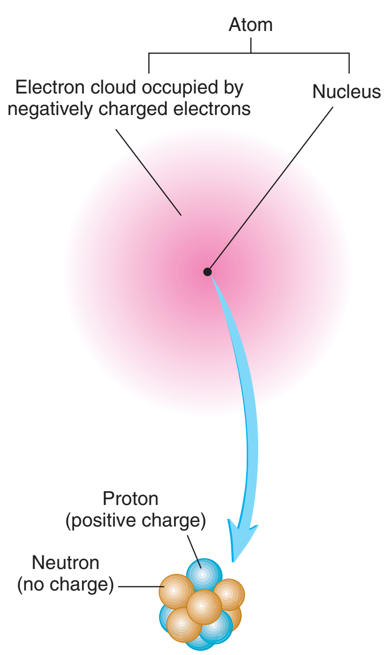



**Figure 2.1** **Model of an Atom**

The tiny, dense nucleus consists of positively charged protons and uncharged neutrons. Most of the volume of an atom is occupied by rapidly moving, negatively charged electrons, which can be represented as an electron cloud. The probable location of an electron is indicated by the color of the electron cloud. The darker the color in each small part of the electron cloud, the more likely the electron is located there.

Protons and neutrons form the **nucleus,** and electrons are moving around the nucleus (see figure 2.1). The nucleus accounts for 99.97% of an atom's mass but only 1 ten-trillionth of its volume. Most of the volume of an atom is occupied by the electrons. Although it is impossible to know precisely where any given electron is located at any particular moment, the region where it is most likely to be found can be represented by an **electron cloud.** The likelihood of locating an electron at a specific point in a region correlates with the darkness of that region in the diagram. The darker the color, the greater the likelihood of finding the electron there at any given moment.

### Atomic Number and Mass Number

The **atomic number** of an element is equal to the number of protons in each atom, and because the number of electrons and protons is equal, the atomic number also indicates the number of electrons. Each element is uniquely defined by the number of protons in the atoms of that element. For example, only hydrogen atoms have one proton, only carbon atoms have six protons, and only oxygen atoms have eight protons (figure 2.2; see table 2.1).

Scientists have been able to create new elements by changing the number of protons in the nuclei of existing elements. Protons, neutrons, or electrons from one atom are accelerated to very high speeds and then smashed into the nucleus of another atom. The resulting changes in the nucleus produce a new element with a new atomic number. To date, 20 elements with an atomic number greater than 92 have been synthesized in this fashion. These artificially produced elements are usually unstable, and they quickly convert back to more stable elements.

Protons and neutrons have about the same mass, and they are responsible for most of the mass of atoms. Electrons, on the other hand, have very little mass. The **mass number** of an element is the number of protons plus the number of neutrons in each atom. For example, the mass number for carbon is 12 because it has six protons and six neutrons.

**P R E D I C T 2**

**The atomic number of potassium is 19, and the mass number is 39. What is the number of protons, neutrons, and electrons in an atom of potassium?**

### Isotopes and Atomic Mass

**Isotopes** (ī′sō-tōpz) are two or more forms of the same element that have the same number of protons and electrons but a different number of neutrons. Thus isotopes have the same atomic number but different mass numbers. There are three isotopes of hydrogen: hydrogen, deuterium, and tritium. All three isotopes have one proton and one electron, but hydrogen has no neutrons in its nucleus, deuterium has one neutron, and tritium has two neutrons (figure 2.3). Isotopes can be denoted using the symbol of the element preceded by the mass number (number of protons and neutrons) of the isotope. Thus hydrogen is $^{1}H$, deuterium is $^{2}H$, and tritium is $^{3}H$.

Individual atoms have very little mass. A hydrogen atom has a mass of $1.67 \times 10^{-24}$ g (see appendix B for an explanation of the scientific notation of numbers). To avoid using such small

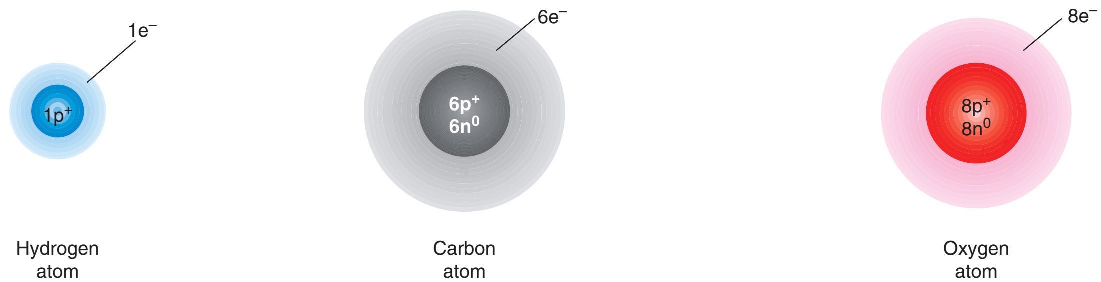



**Figure 2.2** **Hydrogen, Carbon, and Oxygen Atoms**

Within the nucleus, the number of positively charged protons ($p^+$) and uncharged neutrons ($n^0$) is indicated. The negatively charged electrons ($e^-$) are around the nucleus. Atoms are electrically neutral because the number of protons and electrons within an atom are equal.

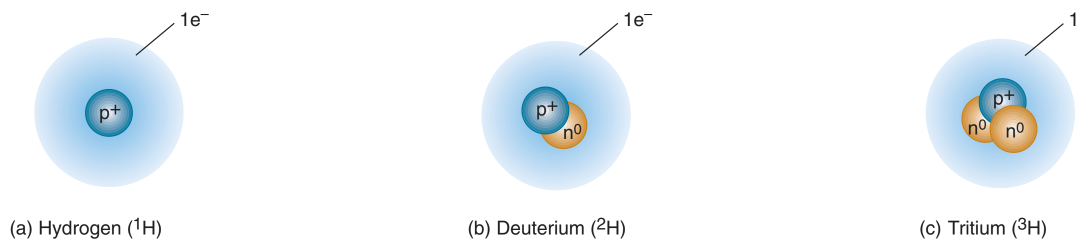



**Figure 2.3** **Isotopes of Hydrogen**

(*a*) Hydrogen has one proton and no neutrons in its nucleus. (*b*) Deuterium has one proton and one neutron in its nucleus. (*c*) Tritium has one proton and two neutrons in its nucleus.

numbers, a system of relative atomic mass is used. In this system, a **unified atomic mass unit (u),** or **dalton (D),** is 1/12 the mass of $^{12}C$, a carbon atom with six protons and six neutrons. Thus $^{12}C$ has an atomic mass of exactly 12 u. A naturally occurring sample of carbon, however, contains mostly $^{12}C$ but also a small quantity of other carbon isotopes such as $^{13}C$, which has six protons and seven neutrons. The **atomic mass** of an element is the *average* mass of its naturally occurring isotopes, taking into account the relative abundance of each isotope. For example, the atomic mass of the element carbon is 12.01 u (see table 2.1), which is slightly more than 12 u because of the additional mass of the small amount of other carbon isotopes. Because the atomic mass is an average, a sample of carbon can be treated as if all the carbon atoms have an atomic mass of 12.01 u.

1. ***Define matter. How is the mass and the matter of an object different?***
2. ***Define element and atom. What four elements are found in the greatest abundance in humans?***
3. ***For each subatomic particle of an atom, state its charge and location. Which subatomic particles are most responsible for the mass and volume of an atom? Which subatomic particles determine atomic number and mass number?***
4. ***Define isotopes and give an example. Define atomic mass. Why is the atomic mass of most elements not exactly equal to the mass number?***

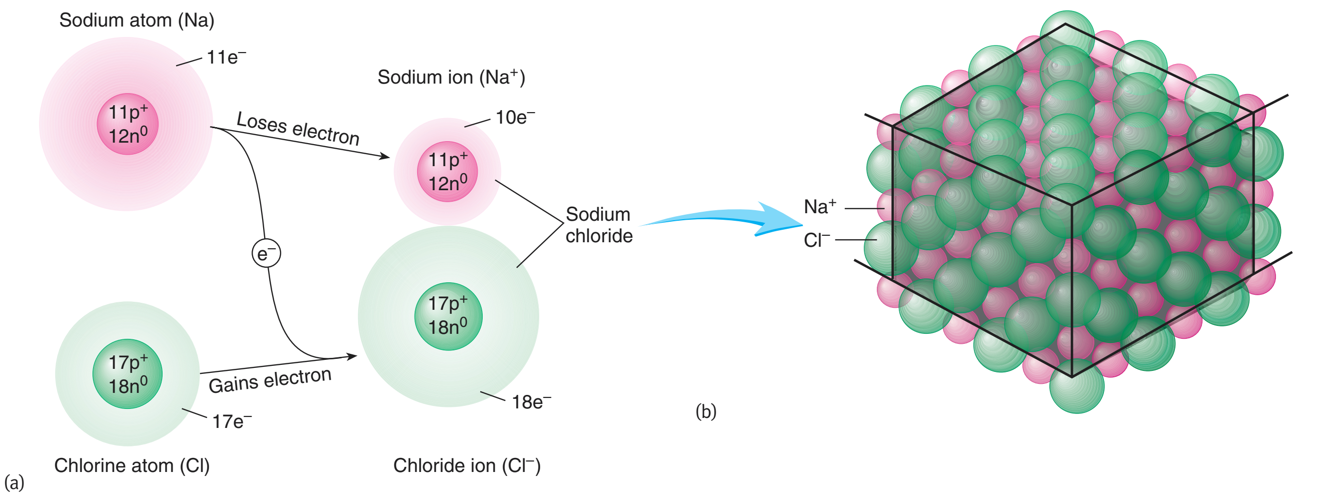
## Electrons and Chemical Bonding

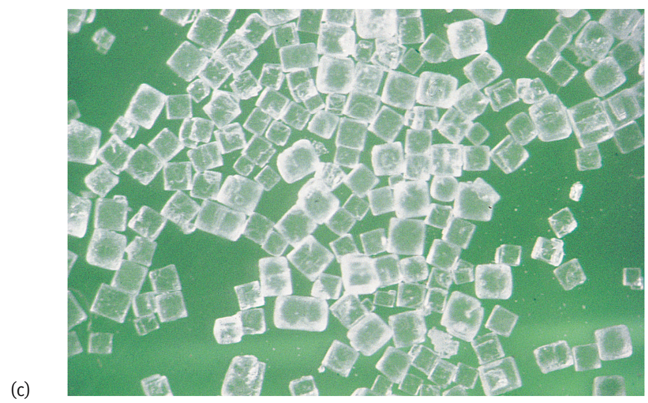
The outermost electrons of an atom determine its chemical behavior. When these outermost electrons are transferred or shared between atoms, **chemical bonding** occurs. Two major types of chemical bonding are ionic and covalent bonding.

### *Ionic Bonding*

An atom is electrically neutral because it has an equal number of protons and electrons. If an atom loses or gains electrons, the number of protons and electrons are no longer equal, and a charged particle called an **ion** (ī′on) is formed. After an atom loses an electron, it has one more proton than it has electrons and is positively charged. A sodium atom (Na) can lose an electron to become a positively charged sodium ion ($Na^+$) (figure 2.4*a*). After an atom gains an electron, it has one more electron than it has protons and is negatively charged. A chlorine atom (Cl) can accept an electron to become a negatively charged chloride ion ($Cl^-$).

Positively charged ions are called **cations** (kat′ī-onz), and negatively charged ions are called **anions** (an′ī-onz). Because oppositely charged ions are attracted to each other, cations and anions tend to remain close together, which is called **ionic** (ī-on′ik) **bonding.** For example, sodium and chloride ions are held together by ionic bonding to form an array of ions called sodium chloride, or table salt (see figure 2.4*b* and *c*). Some ions commonly found in the body are listed in table 2.2.




## Figure 2.4 Ionic Bonding

(*a*) A sodium atom loses an electron to become a smaller-sized positively charged ion, and a chlorine atom gains an electron to become a larger-sized negatively charged ion. The attraction between the oppositely charged ions results in an ionic bond and the formation of sodium chloride. (*b*) The sodium and chlorine ions are organized to form a cube-shaped array. (*c*) Microphotograph of salt crystals reflects the cubic arrangement of the ions.

**Table 2.2 Important Ions**

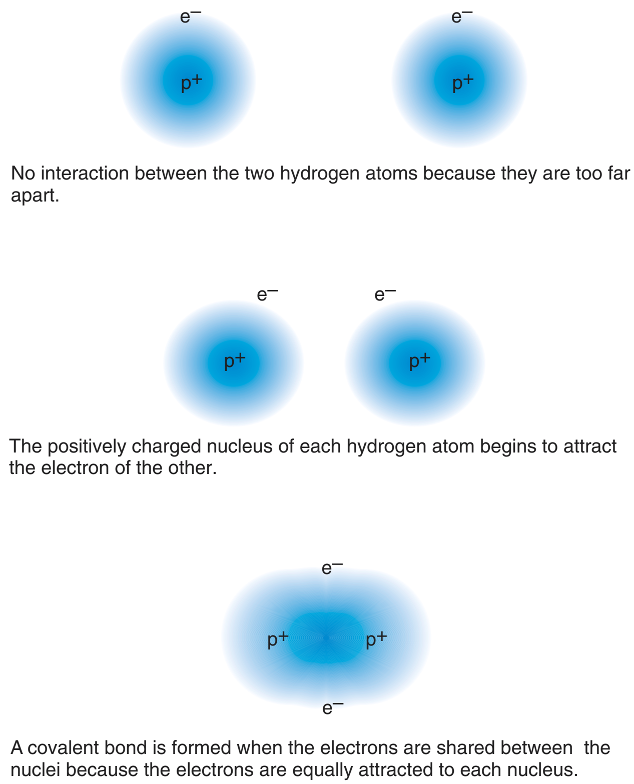
| Common Ions | Symbols | Functions |
|---|---|---|
| Calcium | $Ca^{2+}$ | Bones, teeth, blood clotting, muscle contraction, release of neurotransmitters |
| Sodium | $Na^+$ | Membrane potentials, water balance |
| Potassium | $K^+$ | Membrane potentials |
| Hydrogen | $H^+$ | Acid–base balance |
| Hydroxide | $OH^-$ | Acid–base balance |
| Chloride | $Cl^-$ | Water balance |
| Bicarbonate | $HCO_3^-$ | Acid–base balance |
| Ammonium | $NH_4^+$ | Acid–base balance |
| Phosphate | $PO_4^{3-}$ | Bone, teeth, energy exchange, acid–base balance |
| Iron | $Fe^{2+}$ | Red blood cell formation |
| Magnesium | $Mg^{2+}$ | Necessary for enzymes |
| Iodide | $I^-$ | Present in thyroid hormones |

### *Covalent Bonding*

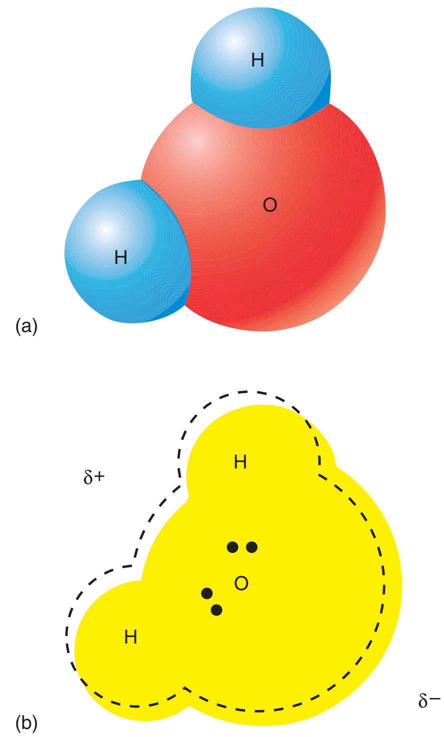
**Covalent bonding** results when atoms share one or more pairs of electrons. The resulting combination of atoms is called a molecule. An example is the covalent bond between two hydrogen atoms to form a hydrogen molecule (figure 2.5). Each hydrogen atom has one electron. As the two hydrogen atoms get closer together, the positively charged nucleus of each atom begins to attract the electron of the other atom. At an optimal distance, the two nuclei mutually attract the two electrons, and each electron is shared by both nuclei. The two hydrogen atoms are now held together by a covalent bond.

When an electron pair is shared between two atoms, a **single covalent bond** results. A single covalent bond is represented by a single line between the symbols of the atoms involved (e.g., H—H). A **double covalent bond** results when two atoms share four electrons, two from each atom. When a carbon atom combines with two oxygen atoms to form carbon dioxide, two double covalent bonds are formed. Double covalent bonds are indicated by a double line between the atoms (O═C═O).

When electrons are shared equally between atoms, as in a hydrogen molecule, the bonds are called **nonpolar covalent bonds.** Atoms bound to one another by a covalent bond do not always share their electrons equally, however, because the nucleus of one atom attracts the electrons more strongly than does the nucleus of




## Figure 2.5 Covalent Bonding




## Figure 2.6 Polar Covalent Bonds

(*a*) A water molecule forms when two hydrogen atoms form covalent bonds with an oxygen atom. (*b*) Electron pairs (*indicated by dots*) are shared between the hydrogen atoms and oxygen. The electrons are shared unequally, as shown by the electron cloud (*yellow*) not coinciding with the dashed outline. Consequently, the oxygen side of the molecule has a slight negative charge (*indicated by* $\delta^-$) and the hydrogen side of the molecule has a slight positive charge (*indicated by* $\delta^+$).

the other atom. Bonds of this type are called **polar covalent bonds** and are common in both living and nonliving matter.

Polar covalent bonds can result in polar molecules, which are electrically asymmetric. For example, oxygen atoms attract electrons more strongly than do hydrogen atoms. When covalent bonding between an oxygen atom and two hydrogen atoms forms a water molecule, the electrons are located in the vicinity of the oxygen nucleus more than in the vicinity of the hydrogen nuclei. Because electrons have a negative charge, the oxygen side of the molecule is slightly more negative than the hydrogen side (figure 2.6).

## Molecules and Compounds

A **molecule** is formed when two or more atoms chemically combine to form a structure that behaves as an independent unit. The atoms that combine to form a molecule can be of the same type, such as two hydrogen atoms combining to form a hydrogen molecule. More typically, a molecule consists of two or more different types of atoms, such as two hydrogen atoms and an oxygen atom forming water. Thus, a glass of water consists of a collection of individual water molecules positioned next to one another.

A **compound** is a substance composed of two or more *different* types of atoms that are chemically combined. Not all molecules are compounds. For example, a hydrogen molecule is not a compound because it does not consist of different types of atoms. Many molecules are compounds, however. Most covalent substances consist of molecules because their atoms form distinct units as a result of the joining of the atoms to each other by a pair of shared electrons. For example, a water molecule is a covalent compound.

On the other hand, ionic compounds are not molecules because the ions are held together by the force of attraction between opposite charges. A piece of sodium chloride does not consist of sodium chloride molecules positioned next to each other. Instead, table salt is an organized array of sodium and chloride ions in which each charged ion is surrounded by several ions of the opposite charge (see figure 2.4*b*). Sodium chloride is an example of a substance that is a compound but is not a molecule.

The kinds and numbers of atoms (or ions) in a molecule or compound can be represented by a formula consisting of the symbols of the atoms (or ions) plus subscripts denoting the number of each type of atom (or ion). The formula for glucose (a sugar) is $C_6H_{12}O_6$, indicating that glucose has 6 carbon, 12 hydrogen, and 6 oxygen atoms (table 2.3).

## Clinical Focus Radioactive Isotopes and X Rays

Protons, neutrons, and electrons are responsible for the chemical properties of atoms. They also have other properties that can be useful in a clinical setting. For example, they have been used to develop methods for examining the inside of the body. **Radioactive isotopes** are commonly used by clinicians and researchers because sensitive measuring devices can detect their radioactivity, even when they are present in very small amounts.

Radioactive isotopes have unstable nuclei that spontaneously change to form more stable nuclei. As a result, either new isotopes or new elements are produced. In this process of nuclear change, alpha particles, beta particles, and gamma rays are emitted from the nuclei of radioactive isotopes. Alpha ($\alpha$) particles are positively charged helium ions ($He^{2+}$), which consist of two protons and two neutrons. Beta ($\beta$) particles are electrons formed as neutrons change into protons. An electron is ejected from the neutron, and the proton that is produced remains in the nucleus. Gamma ($\gamma$) rays are a form of electromagnetic radiation (high-energy photons) released from nuclei as they lose energy.

All isotopes of an element have the same atomic number, and their chemical behavior is very similar. For example, $^{3}H$ (tritium) can substitute for $^{1}H$ (hydrogen), and either $^{125}$iodine or $^{131}$iodine can substitute for $^{126}$iodine in chemical reactions.

Several procedures that are used to determine the concentration of substances such as hormones depend on the incorporation of small amounts of radioactive isotopes, such as $^{125}$iodine, into the substances being measured. Clinicians using these procedures can more accurately diagnose disorders of the thyroid gland, the adrenal gland, and the reproductive organs.

Radioactive isotopes are also used to treat cancer. Some of the particles released from isotopes have a very high energy content and can penetrate and destroy tissues. Thus radioactive isotopes can be used to destroy tumors because rapidly growing tissues such as tumors are more sensitive to radiation than healthy cells. Radiation can also be used to sterilize materials that cannot be exposed to high temperatures (e.g., some fabric and plastic items used during surgical procedures). In addition, radioactive emissions can be used to sterilize food and other items.

X rays are electromagnetic radiations with a much shorter wavelength than visible light. When electric current is used to heat a filament to very high temperatures, energy of the electrons becomes so great that some electrons are emitted from the hot filament. When these electrons strike a positive electrode at high speeds, they release some of their energy in the form of x rays.

X rays do not penetrate dense material as readily as they penetrate less dense material, and x rays can expose photographic film. Consequently, an x-ray beam can pass through a person and onto photographic film. Dense tissues of the body absorb the x rays, and in these areas the film is underexposed and so appears white or light in color on the developed film. On the other hand, the x rays readily pass through less dense tissue, and the film in these areas is overexposed and appears black or dark in color. In an x-ray film of the skeletal system the dense bones are white, and the less dense soft tissues are dark, often so dark that no details can be seen. Because the dense bone material is clearly visible, x rays can be used to determine whether bones are broken or have other abnormalities.

Soft tissues can be photographed by using low-energy x rays. Mammograms are low-energy x rays of the breast that can be used to detect tumors, because tumors are slightly denser than normal tissue.

Radiopaque substances are dense materials that absorb x rays. If a radiopaque liquid is given to a patient, the liquid assumes the shape of the organ into which it is placed. For example, if a barium solution is swallowed, the outline of the upper digestive tract can be photographed using x rays to detect such abnormalities as ulcers.

The **molecular mass** of a molecule or compound can be determined by adding up the atomic masses of its atoms (or ions). The term *molecular mass* is used for convenience for ionic compounds, even though they are not molecules. For example, the atomic mass of sodium is 22.99 and chloride is 35.45. The molecular mass of NaCl is therefore 58.44 (22.99 + 35.45).

*5. Describe how ionic bonding occurs. What is a cation and an anion?*
*6. Describe how covalent bonding occurs. What is the difference between polar and nonpolar covalent bonds?*
*7. Distinguish between a molecule and a compound. Are all molecules compounds? Are all compounds molecules?*
*8. Define molecular mass.*

**P R E D I C T 3**

**What is the molecular mass of a molecule of glucose? (Use table 2.1.)**

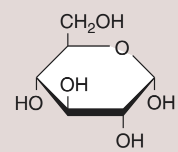
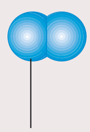
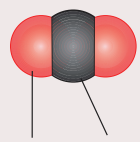
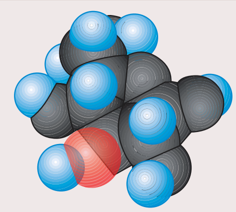
## Intermolecular Forces

**Intermolecular forces** result from the weak electrostatic attractions between the oppositely charged parts of molecules, or between ions and molecules. Intermolecular forces are much weaker than the forces producing chemical bonding.

### *Hydrogen Bonds*

Molecules with polar covalent bonds have positive and negative “ends.” Intermolecular force results from the attraction of the positive end of one polar molecule to the negative end of another

**Table 2.3 Picturing Molecules**

| Representation | Hydrogen | Carbon Dioxide | Glucose |
|---|---|---|---|
| **Chemical Formula**<br>Shows the kind and number of atoms present. | $H^2$ | $CO_2$ | $C_6H_{12}O_6$ |
| **Electron-Dot Formula**<br>The bonding electrons are shown as dots between the symbols of the atoms. | H:H<br>Single covalent bond | O::C::O<br>Double covalent bond | Not used for complex molecules |
| **Bond-Line Formula**<br>The bonding electrons are shown as lines between the symbols of the atoms. | H—H<br>Single covalent bond | O═C═O<br>Double covalent bond |  |
| **Models**<br>Atoms are shown as different-sized and different-colored spheres. |  |  |  |

polar molecule. When hydrogen forms a covalent bond with oxygen, nitrogen, or fluorine, the resulting molecule becomes very polarized. If the positively charged hydrogen of one molecule is attracted to the negatively charged oxygen, nitrogen, or fluorine of another molecule, a **hydrogen bond** is formed. For example, the positively charged hydrogen atoms of a water molecule form hydrogen bonds with the negatively charged oxygen atoms of other water molecules (figure 2.7).

Hydrogen bonds play an important role in determining the shape of complex molecules because the hydrogen bonds between different polar parts of the molecule hold the molecule in its normal three-dimensional shape (see the sections "Proteins" and "Nucleic Acids: DNA and RNA" later in this chapter).

Table 2.4 summarizes the important characteristics of chemical bonding (ionic and covalent) and intermolecular forces (hydrogen bonds).

### *Solubility and Dissociation*

**Solubility** is the ability of one substance to dissolve in another, for example, when sugar dissolves in water. Charged substances such as sodium chloride, and polar substances such as glucose, dissolve

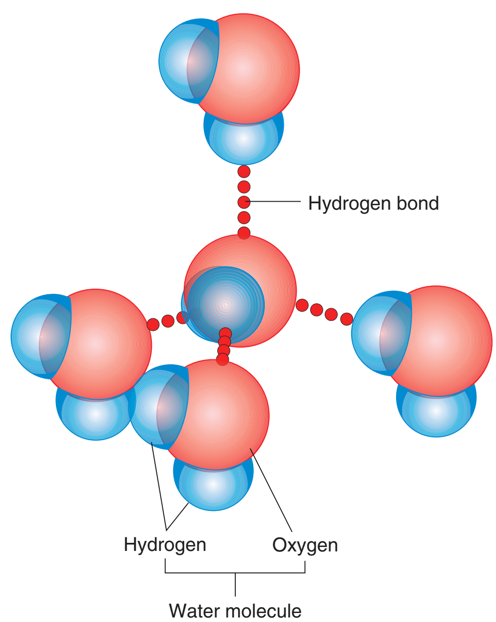
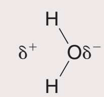
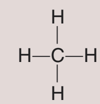
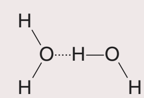



**Figure 2.7 Hydrogen Bonds**

The positive hydrogen part of one water molecule forms a hydrogen bond (*red dotted line*) with the negative oxygen part of another water molecule. As a result, hydrogen bonds hold the water molecules together.

**Table 2.4 Comparison of Bonds**

| Definition | Charge Distribution | Example |
|---|---|---|
| **Ionic Bond** | | |
| Complete transfer of electrons between two atoms | Separate positively charged and negatively charged ions | $Na^+Cl^-$ Sodium chloride |
| **Polar Covalent Bond** | | |
| Unequal sharing of electrons between two atoms | Slight positive charge ($\delta^+$) on one side of the molecule and slight negative charge ($\delta^-$) on the other side of the molecule |  Water |
| **Nonpolar Covalent Bond** | | |
| Equal sharing of electrons between two atoms | Charge evenly distributed among the atoms of the molecule |  Methane |
| **Hydrogen Bond** | | |
| Attraction of oppositely charged ends of one polar molecule to another polar molecule | Charge distribution within the polar molecules results from polar covalent bonds |  Water molecules |

in water readily, whereas nonpolar substances such as oils do not. We all have seen how oil floats on water. Substances dissolve in water when they become surrounded by water molecules. If the positive and negative ends of the water molecules are attracted more to the charged ends of other molecules than they are to each other, the hydrogen bonds between the ends of the water molecules are broken, and the water molecules surround the other molecules, which become dissolved in the water.

When ionic compounds dissolve in water, their ions **dissociate,** or separate, from one another because the cations are attracted to the negative ends of the water molecules, and the anions are attracted to the positive ends of the water molecules. When sodium chloride dissociates in water, the sodium and chloride ions separate, and water molecules surround and isolate the ions, thereby keeping them in solution (figure 2.8).

When molecules (covalent compounds) dissolve in water, they usually remain intact even though they are surrounded by water molecules. Thus, in a glucose solution, glucose molecules are surrounded by water molecules.

Cations and anions that dissociate in water are sometimes called **electrolytes** (ē-lek′trō-lītz) because they have the capacity to conduct an electric current, which is the flow of charged particles. An electrocardiogram (ECG) is a recording of electric currents produced by the heart. These currents can be detected by electrodes on the surface of the body because the ions in the body fluids conduct electric currents. Molecules that do not dissociate form solutions that do not conduct electricity and are called **nonelectrolytes.**

*9. Define hydrogen bond, and explain how hydrogen bonds hold polar molecules, such as water, together. How do hydrogen bonds affect the shape of a molecule?*

*10. Define solubility. How do ionic and covalent compounds typically dissolve in water?*

*11. Distinguish between electrolytes and nonelectrolytes.*

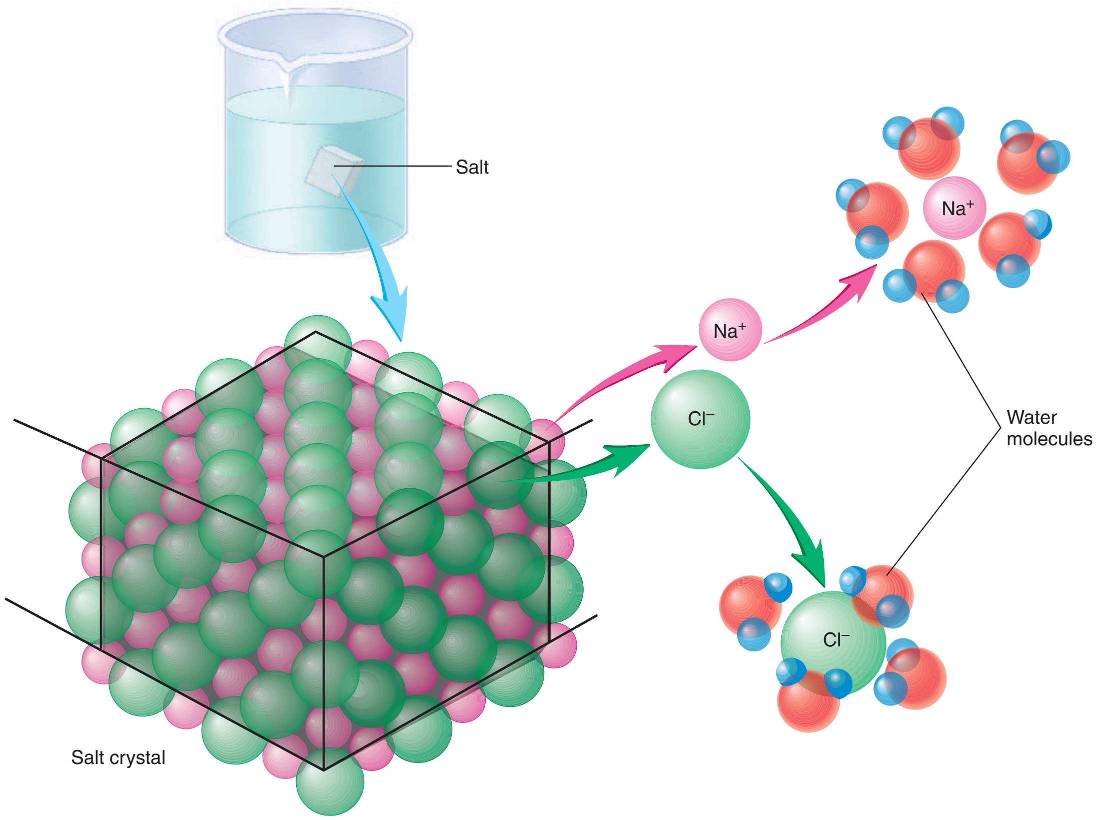
## Chemical Reactions and Energy

### *Objectives*

- *Describe and give examples of the types of chemical reactions occurring in the body.*
- *Define potential and kinetic energy. Describe mechanical, chemical, and heat energy as they relate to the human body.*
- *List the factors that affect the speed of a chemical reaction.*

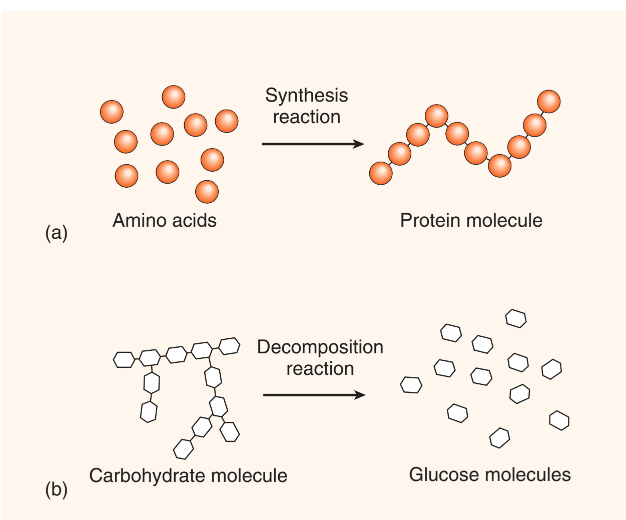
In a **chemical reaction,** atoms, ions, molecules, or compounds interact either to form or to break chemical bonds. The substances that enter into a chemical reaction are called the **reactants,** and the substances that result from the chemical reaction are called the **products.**

For our purposes, three important points can be made about chemical reactions. First, in some reactions, less complex reactants are combined to form a larger, more complex product. An example is the synthesis of the complex molecules of the human body from basic "building blocks" obtained in food (figure 2.9*a*). Second, in other reactions, a reactant can be broken down, or decomposed, into simpler, less complex products. An example is the breakdown of food molecules into basic building blocks. (figure 2.9*b*). Third, atoms are generally associated with other atoms through chemical bonding or intermolecular forces; therefore, to synthesize new products or break down reactants it is necessary to change the relationship between atoms.




## Figure 2.8 Dissociation

Sodium chloride (table salt) dissociating in water. The positively charged sodium ions ($Na^+$) are attracted to the negative oxygen (*red*) end of the water molecule, and the negatively charged chlorine ions ($Cl^-$) are attracted to the positively charged hydrogen (*blue*) end of the water molecule.




## Figure 2.9 Synthesis and Decomposition Reactions

(*a*) Synthesis reaction in which amino acids, the basic "building blocks" of proteins, combine to form a protein molecule. (*b*) Decomposition reaction in which a complex carbohydrate breaks down into smaller glucose molecules, which are the "building blocks" of carbohydrates.

## Synthesis Reactions

When two or more reactants chemically combine to form a new and larger product, the process is called a **synthesis reaction.** An example of a synthesis reaction is the combination of two amino acids to form a dipeptide (figure 2.10*a*). In this particular synthesis reaction, water is removed from the amino acids as they are bound together. Synthesis reactions in which water is a product are called **dehydration** (water out) **reactions.** Note that old chemical bonds are broken and new chemical bonds are formed as the atoms rearrange as a result of a synthesis reaction.

Another example of a synthesis reaction in the body is the formation of adenosine triphosphate (ATP). In ATP, *A* stands for adenosine, *T* stands for tri- or three, and *P* stands for phosphate group ($PO_4^{3-}$). Thus, ATP consists of adenosine and three phosphate groups (see p. 53 for the details of the structure of ATP). ATP is synthesized from adenosine diphosphate (ADP), which has two phosphate groups, and an inorganic phosphate ($H_2PO_4^-$), which is often symbolized as $P_i$.

| A-P-P | + | $P_i$ | → | A-P-P-P |
|---|---|---|---|---|
| (ADP) | | (Inorganic phosphate) | | (ATP) |

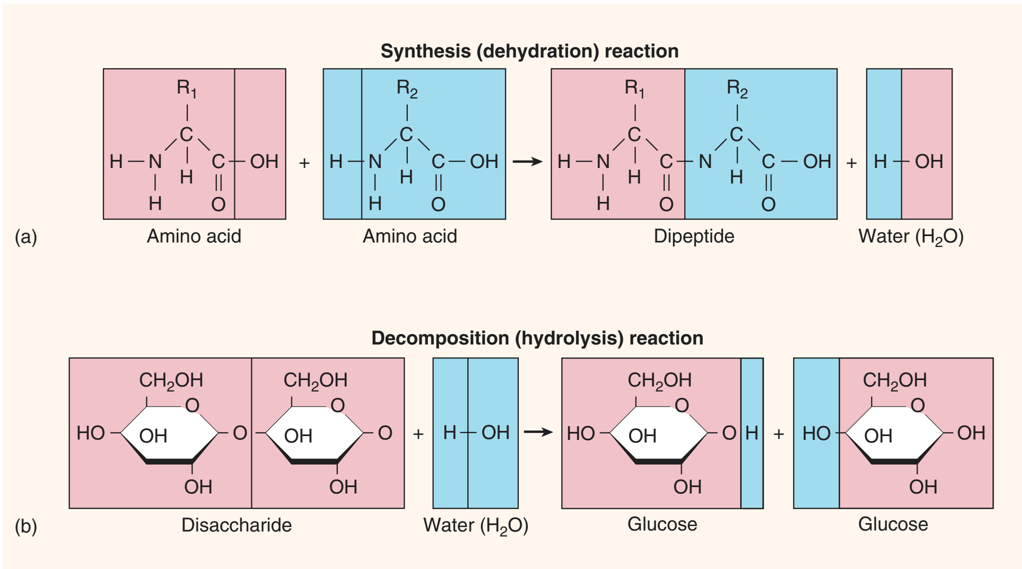



**Figure 2.10** **Synthesis (Dehydration) and Decomposition (Hydrolysis) Reactions**

(*a*) Synthesis reaction in which two amino acids combine to form a dipeptide. This reaction is also a dehydration reaction because it results in the removal of a water molecule from the amino acids. (*b*) Decomposition reaction in which a disaccharide breaks apart to form glucose molecules. This reaction is also a hydrolysis reaction because it involves the splitting of a water molecule.

Synthesis reactions produce the molecules characteristic of life, such as ATP, proteins, carbohydrates, lipids, and nucleic acids. All of the synthesis reactions that occur within the body are referred to collectively as **anabolism** (ă-nab′ō-lizm). The growth, maintenance, and repair of the body could not take place without anabolic reactions.

## Decomposition Reactions

The term *decompose* means to break down into smaller parts. A **decomposition reaction** is the reverse of a synthesis reaction—a larger reactant is chemically broken down into two or more smaller products. The breakdown of a disaccharide (a type of carbohydrate) into glucose molecules (figure 2.10*b*) is an example. Note that this particular reaction requires that water be split into two parts and that each part be contributed to one of the new glucose molecules. Reactions that use water in this manner are called **hydrolysis** (hī-drol′i-sis; water dissolution) **reactions.**

The breakdown of ATP to ADP and an inorganic phosphate is another example of a decomposition reaction.

| | | | | |
|---|---|---|---|---|
| A-P-P-P | → | A-P-P | + | $P_i$ |
| (ATP) | | (ADP) | | (Inorganic phosphate) |

The decomposition reactions that occur in the body are collectively called **catabolism** (kă-tab′ō-lizm). They include the digestion of food molecules in the intestine and within cells, the breakdown of fat stores, and the breakdown of foreign matter and microorganisms in certain blood cells that function to protect the body. All of the anabolic and catabolic reactions in the body are collectively defined as **metabolism.**

## Reversible Reactions

A **reversible reaction** is a chemical reaction in which the reaction can proceed from reactants to products or from products to reactants. When the rate of product formation is equal to the rate of the reverse reaction, the reaction system is said to be at **equilibrium.** At equilibrium the amount of reactants relative to the amount of products remains constant.

The following analogy may help to clarify the concept of reversible reactions and equilibrium. Imagine a trough containing water. The trough is divided into two compartments by a partition, but the partition contains holes that allow water to move freely between the compartments. Because water can move in either direction, this is like a reversible reaction. Let the water in the left compartment be the reactant and the water in the right compartment be the product. At equilibrium, the amount of reactant relative to the amount of product in each compartment is always the same because the partition allows water to pass between the two compartments until the level of water is the same in both compartments. If additional water is added to the reactant compartment, water flows from it through the partition to the product compartment until the level of water is the same in both compartments. Likewise, if additional reactants are added to a reaction system, some will form product until equilibrium is reestablished. Unlike this analogy, however, the amount of the reactants compared to the amount of products of most reversible reactions is not one to one.

Depending on the specific reversible reaction, one part reactant to two parts product, two parts reactant to one part product, or many other possibilities can occur.

An important reversible reaction in the human body involves carbon dioxide and hydrogen ions. The reaction between carbon dioxide ($CO_2$) and water ($H_2O$) to form carbonic acid ($H_2CO_3$) is reversible. Carbonic acid then separates by a reversible reaction to form hydrogen ions ($H^+$) and bicarbonate ions ($HCO_3^-$).

$$CO_2 + H_2O \rightleftarrows H_2CO_3 \rightleftarrows H^+ + HCO_3^+$$

If $CO_2$ is added to $H_2O$, additional $H_2CO_3$ forms, which, in turn, causes more $H^+$ and $HCO_3^-$ to form. The amount of $H^+$ and $HCO_3^-$ relative to $CO_2$ therefore remains constant. Maintaining a constant level of $H^+$ is necessary for proper functioning of the nervous system. This can be achieved, in part, by regulating blood $CO_2$ levels. For example, slowing down the respiration rate causes blood carbon dioxide levels to increase.

**P R E D I C T 4**

**If the respiration rate increases, $CO_2$ is eliminated from the blood. What effect does this change have on blood $H^+$ ion levels?**

## Oxidation–Reduction Reactions

Chemical reactions that result from the exchange of electrons between the reactants are called oxidation–reduction reactions. When sodium and chlorine react to form sodium chloride, the sodium atom loses an electron, and the chlorine atom gains an electron. The loss of an electron by an atom is called **oxidation,** and the gain of an electron is called **reduction.** The transfer of the electron can be complete, resulting in an ionic bond, or it can be a partial transfer, resulting in a covalent bond. Because the complete or partial loss of an electron by one atom is accompanied by the gain of that electron by another atom, these reactions are called **oxidation–reduction reactions.** Synthesis and decomposition reactions can be oxidation–reduction reactions. Thus, it is possible for a chemical reaction to be described in more than one way.

12. ***Define a chemical reaction and compare synthesis and decomposition reactions. How do anabolism, catabolism, and metabolism relate to synthesis and decomposition reactions?***
13. ***Describe a dehydration and a hydrolysis reaction.***
14. ***Describe reversible reactions. What is meant by the equilibrium condition in reversible reactions?***
15. ***What is an oxidation–reduction reaction?***

**P R E D I C T 5**

**When hydrogen gas combines with oxygen gas to form water, is the hydrogen reduced or oxidized? Explain.**

## Energy

Energy, unlike matter, does not occupy space, and it has no mass. **Energy** is defined as the capacity to do **work,** that is, to move matter. Energy can be subdivided into potential energy and kinetic energy. **Potential energy** is stored energy that could do work but is not doing so. **Kinetic** (ki-net′ik) **energy** is the form of energy that actually does work and moves matter. A ball held at arm's length above the floor has potential energy. No energy is expended as long as the ball does not move. If the ball is released and falls toward the floor, however, it has kinetic energy.

According to the conservation of energy principle, energy is neither created nor destroyed. Potential energy, however, can be converted into kinetic energy, and kinetic energy can be converted into potential energy. For example, the potential energy in the ball is converted into kinetic energy as the ball falls toward the floor. Conversely, the kinetic energy required to raise the ball from the floor is converted into potential energy.

Potential and kinetic energy can be found in many different forms. **Mechanical energy** is energy resulting from the position or movement of objects. Many of the activities of the human body, such as moving a limb, breathing, or circulating blood involve mechanical energy. Other forms of energy are chemical energy, heat energy, electric energy, and electromagnetic (radiant) energy.

### *Chemical Energy*

The **chemical energy** of a substance is a form of stored (potential) energy within its chemical bonds. In any given chemical reaction, the potential energy contained in the chemical bonds of the reactants can be compared to the potential energy in the chemical bonds of the products. If the potential energy in the chemical bonds of the reactants is less than that of the products, then energy must be supplied for the reaction to occur. For example, the synthesis of ATP from ADP.

| $ADP + H_2PO_4^-$ | + | Energy | $\rightarrow$ | $ATP + H_2O$ |
|---|---|---|---|---|
| (Less potential energy in reactants) | | | | (More potential energy in products) |

For simplicity, the $H_2O$ is often not shown in this reaction, and $P_i$ is used to represent inorganic phosphate ($H_2PO_4^-$). For this reaction to occur, bonds in $H_2PO_4^-$ are broken and bonds are formed in ATP and $H_2O$. As a result of the breaking of existing bonds, the formation of new bonds, and the input of energy, these products have more potential energy than the reactants (figure 2.11*a*).

If the potential energy in the chemical bonds of the reactants is greater than that of the products, energy is released by the reaction. For example, the chemical bonds of food molecules contain more potential energy than the waste products that are produced when food molecules are decomposed. The energy released from the chemical bonds of food molecules is used by living systems to synthesize ATP. Once ATP is produced, the breakdown of ATP to ADP results in the release of energy.

| $ATP + H_2O$ | $\rightarrow$ | $ADP + H_2PO_4^-$ | + | Energy |
|---|---|---|---|---|
| (More potential energy in reactants) | | (Less potential energy in products) | | |

For this reaction to occur, the bonds in ATP and $H_2O$ are broken and bonds in $H_2PO_4^-$ are formed. As a result of breaking the existing bonds and forming new bonds, these products have less potential energy than the reactants, and energy is released

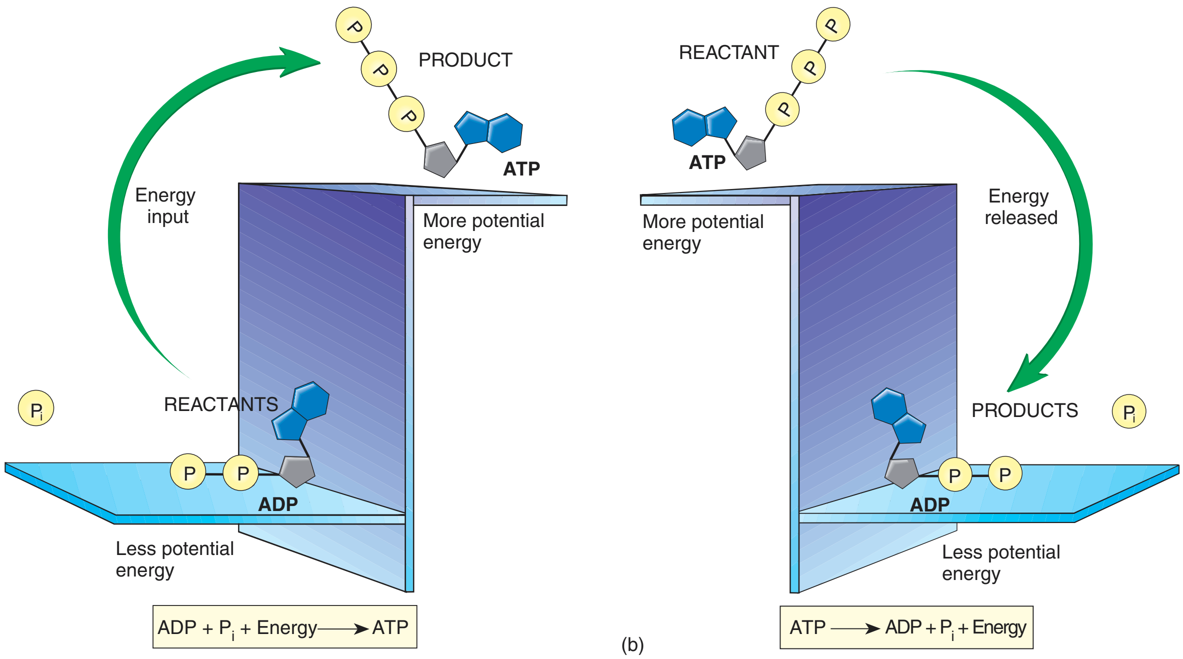



**Figure 2.11** **Energy and Chemical Reactions**

In each figure the upper shelf represents a higher energy level, and the lower shelf represents a lower energy level. (*a*) Reaction in which the input of energy is required for the synthesis of ATP. (*b*) Reaction in which energy is released as a result of the breakdown of ATP.

(figure 2.11*b*). Note that the energy released does not come from breaking the phosphate bond of ATP, because breaking a chemical bond requires the input of energy. It is commonly stated, however, that the breakdown of ATP results in the release of energy, which is true when the overall reaction is considered. The energy released when ATP is broken down can be used in the synthesis of other molecules; to do work, such as muscle contraction; or to produce heat.

## Heat Energy

**Heat** is the energy that flows between objects that are at different temperatures. For example, when you touch someone who has a fever, you can feel the increased heat from the person's body. Temperature is a measure of how hot or cold a substance is relative to another substance. Heat is always transferred from a hotter object to a cooler object, such as from a hot stove top to a finger.

All other forms of energy can be converted into heat energy. For example, when a moving object comes to rest, its kinetic energy is converted into heat energy by friction. Some of the potential energy of chemical bonds is released as heat energy during chemical reactions. The body temperature of humans is maintained by heat produced in this fashion.

16. ***How is energy different from matter? How are potential and kinetic energy different from each other?***
17. ***Define mechanical energy, chemical energy, and heat energy. How is chemical energy converted to mechanical energy and heat energy in the body?***
18. ***Use ATP and ADP to illustrate the release or input of energy in chemical reactions.***

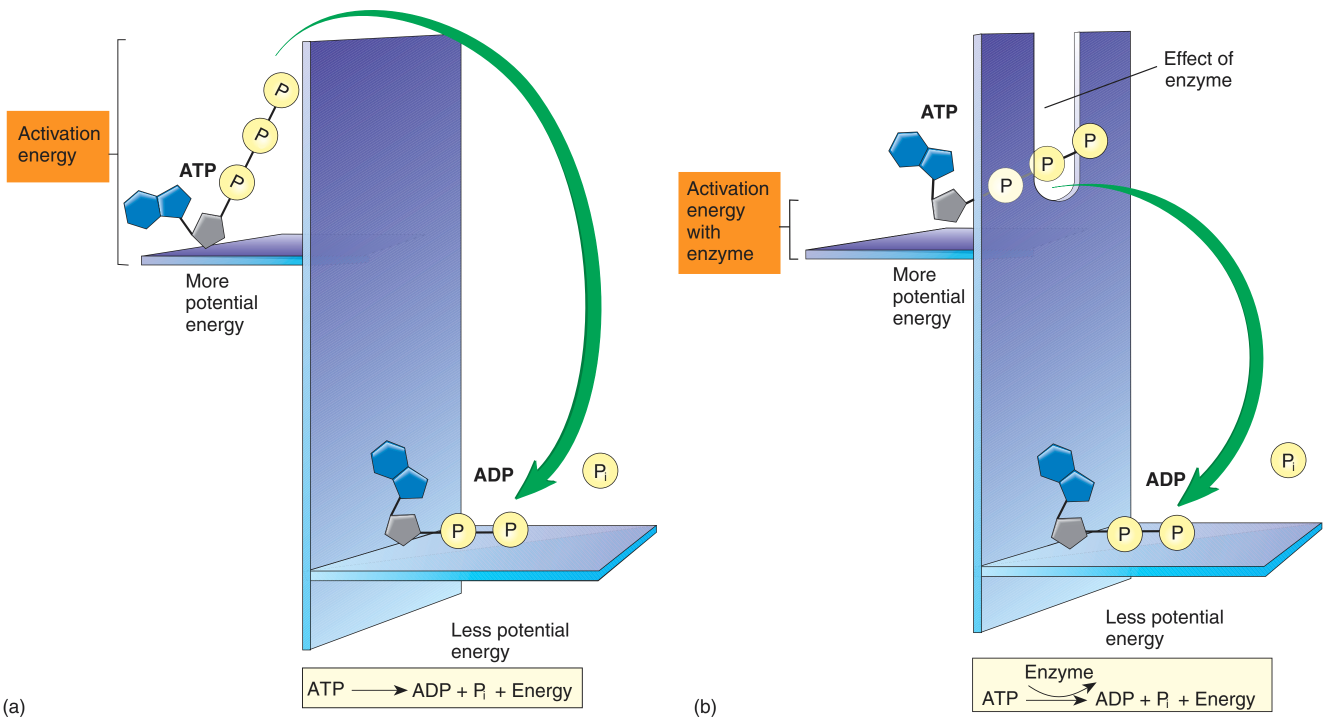
**P R E D I C T 6**

**Energy from the breakdown of ATP provides the kinetic energy for muscle movement. Why does body temperature increase during exercise?**

## Speed of Chemical Reactions

Molecules are constantly in motion and therefore have kinetic energy. A chemical reaction occurs only when molecules with sufficient kinetic energy collide with each other. As two molecules move closer together, the negatively charged electron cloud of one molecule repels the negatively charged electron cloud of the other molecule. If the molecules have sufficient kinetic energy, they overcome this repulsion and come together. The nuclei in some atoms attract the electrons of other atoms, resulting in the breaking and formation of new chemical bonds. The **activation energy** is the minimum energy that the reactants must have to start a chemical reaction (figure 2.12*a*). Even reactions that result in a release of energy must overcome the activation energy barrier for the reaction to proceed. For example, heat in the form of a spark is required to start the reaction between oxygen and gasoline vapor. Once some oxygen molecules react with gasoline, the energy released can start additional reactions.

Given any population of molecules, some of them have more kinetic energy and move about faster than others. Even so, at normal body temperatures, most of the chemical reactions necessary




## Figure 2.12 Activation Energy and Enzymes

(*a*) Activation energy is needed to change ATP to ADP. The upper shelf represents a higher energy level, and the lower shelf represents a lower energy level. The "wall" extending above the upper shelf represents the activation energy. Even though energy is given up moving from the upper to the lower shelf, the activation energy "wall" must be overcome before the reaction can proceed. (*b*) The enzyme lowers the activation energy, making it easier for the reaction to proceed.

for life proceed too slowly to support life because few molecules have enough energy to start a chemical reaction. **Catalysts** (kat′ă-listz) are substances that increase the rate of chemical reactions without being permanently changed or depleted. **Enzymes** (en′zīmz), which are discussed in greater detail on p. 49, are protein catalysts. Enzymes increase the rate of chemical reactions by lowering the activation energy necessary for the reaction to begin (figure 2.12*b*). As a result, more molecules have sufficient energy to undergo chemical reactions. With an enzyme, the rate of a chemical reaction can take place more than a million times faster than without the enzyme.

Temperature can also affect the speed of chemical reactions. As temperature increases, reactants have more kinetic energy, move at faster speeds, and collide with one another more frequently and with greater force, thereby increasing the likelihood of a chemical reaction. When a person has a fever of only a few degrees, reactions occur throughout the body at an accelerated rate, resulting in increased activity in the organ systems such as increased heart and respiratory rates. When body temperature drops, various metabolic processes slow. In cold weather, the fingers are less agile largely because of the reduced rate of chemical reactions in cold muscle tissue.

Within limits, the greater the concentration of the reactants, the greater the rate at which a given chemical reaction proceeds. This occurs because, as the concentration of reactants increases, they are more likely to come into contact with one another. For example, the normal concentration of oxygen inside cells enables oxygen to come into contact with other molecules and produce the chemical reactions necessary for life. If the oxygen concentration decreases, the rate of chemical reactions decreases. This decrease can impair cell function and even result in death.

***19. Define activation energy, catalysts, and enzymes. How do enzymes increase the rate of chemical reactions?***

***20. What effect does increasing temperature or increasing concentration of the reactants have on the rate of a chemical reaction?***

# Inorganic Chemistry

### *Objectives*

- ***Describe the properties of water that make it important for living organisms.***
- ***Discuss mixtures.***
- ***Define acids, bases, salts, and buffers, and describe the pH scale.***
- ***Explain the importance of oxygen and carbon dioxide to living organisms.***

It was once believed that inorganic substances were those that came from nonliving sources and organic substances were those extracted from living organisms. As the science of chemistry developed, however, it became apparent that organic substances could be manufactured in the laboratory. As defined currently, **inorganic chemistry** generally deals with those substances that do not contain carbon, whereas **organic chemistry** is the study of carbon-containing substances. These definitions have a few exceptions. For example, carbon monoxide (CO), carbon dioxide ($CO_2$), and bicarbonate ion ($HCO_3^-$) are classified as inorganic molecules.

## Water

A molecule of **water** is composed of one atom of oxygen joined to two atoms of hydrogen by covalent bonds. Water molecules are polar, with a partial positive charge associated with the hydrogen atoms and a partial negative charge associated with the oxygen atom. Hydrogen bonds form between the positively charged hydrogen atoms of one water molecule and the negatively charged oxygen atoms of another water molecule. These hydrogen bonds organize the water molecules into a lattice that holds the water molecules together (see figures 2.6 and 2.7).

Water accounts for approximately 50% of the weight of a young adult female and 60% of a young adult male. Females have a lower percentage of water than males because they typically have more body fat, which is relatively free of water. Plasma, the liquid portion of blood, is 92% water. Water has physical and chemical properties well suited for its many functions in living organisms. These properties are outlined in the following sections.

### *Stabilizing Body Temperature*

Water has a high **specific heat,** meaning that a relatively large amount of heat is required to raise its temperature; therefore, it tends to resist large temperature fluctuations. When water evaporates, it changes from a liquid to a gas, and because heat is required for that process, the evaporation of water from the surface of the body rids the body of excess heat.

### *Protection*

Water is an effective lubricant that provides protection against damage resulting from friction. For example, tears protect the surface of the eye from the rubbing of the eyelids. Water also forms a fluid cushion around organs that helps to protect them from trauma. The cerebrospinal fluid that surrounds the brain is an example.

### *Chemical Reactions*

Many of the chemical reactions necessary for life do not take place unless the reacting molecules are dissolved in water. For example, sodium chloride must dissociate in water into sodium and chloride ions before they can react with other ions. Water also directly participates in many chemical reactions. As previously mentioned, a dehydration reaction is a synthesis reaction in which water is produced, and a hydrolysis reaction is a decomposition reaction that requires a water molecule (see figure 2.10).

### *Mixing Medium*

A **mixture** is a combination of two or more substances physically blended together, but not chemically combined. A **solution** is any liquid, gas, or solid in which the substances are uniformly distributed with no clear boundary between the substances. For example, a salt solution consists of salt dissolved in water, air is a solution containing a variety of gases, and wax is a solid solution of several fatty substances. Solutions are often described in terms of one substance dissolving in another: the **solute** (sol′ūt) dissolves in the **solvent.** In a salt solution, water is the solvent and the dissolved salt is the solute. Sweat is a salt solution in which sodium chloride and other solutes are dissolved in water.

A **suspension** is a mixture containing materials that separate from each other unless they are continually, physically blended together. Blood is a suspension containing red blood cells suspended in a liquid called plasma. As long as the red blood cells and plasma are mixed together as they pass through blood vessels, the red blood cells remain suspended in the plasma. If the blood is allowed to sit in a container, however, the red blood cells and plasma separate from each other.

A **colloid** (kol′oyd) is a mixture in which a dispersed (solutelike) substance is distributed throughout a dispersing (solventlike) substance. The dispersed particles are larger than a simple molecule but small enough that they remain dispersed and do not settle out. Proteins, which are large molecules, and water form colloids. For instance, the plasma portion of blood and the liquid interior of cells are colloids containing many important proteins.

In living organisms the complex fluids inside and outside cells consist of solutions, suspensions, and colloids. Blood is an example of all of these mixtures. It is a solution containing dissolved nutrients such as sugar, a suspension holding red blood cells, and a colloid containing proteins.

The ability of water to mix with other substances enables it to act as a medium for transport, moving substances from one part of the body to another. Body fluids such as plasma transport nutrients, gases, waste products, and a variety of molecules involved with regulating body functions.

## Solution Concentrations

The concentration of solute particles dissolved in solvents can be expressed in several ways. One common way is to indicate the percent of solute by weight per volume of solution. A 10% solution of sodium chloride can be made by dissolving 10 g of sodium chloride into enough water to make 100 mL of solution.

Physiologists often determine concentrations in **osmoles** (os′mōlz), which express the number of particles in a solution. A particle can be an atom, ion, or molecule. An osmole (osm) is $6.022 \times 10^{23}$ particles of a substance in 1 kilogram (kg) of water. Just as a grocer sells eggs in lots of 12 (a dozen), a chemist groups atoms in lots of $6.022 \times 10^{23}$. The **osmolality** (os-mō-lal′i-tē) of a solution is a reflection of the number, not the type, of particles in a solution. For example, a 1 osm glucose solution and a 1 osm sodium chloride solution both contain $6.022 \times 10^{23}$ particles per kg water. The glucose solution, however, has $6.022 \times 10^{23}$ molecules of glucose, whereas the sodium chloride dissociates into $3.011 \times 10^{23}$ sodium ions and $3.011 \times 10^{23}$ chloride ions.

Because the concentration of particles in body fluids is so low, the measurement **milliosmole** (mOsm), 1/1000 of an osmole, is used. Most body fluids have a concentration of about 300 mOsm and contain many different ions and molecules. The concentration of body fluids is important because it influences the movement of water into or out of cells (see chapter 3). Appendix C contains more information on calculating concentrations.

***21. Define inorganic and organic chemistry.***
***22. List four functions that water performs in living organisms and give an example of each.***
***23. Describe solutions, suspensions, and colloids, and give an example of each. Define solvent and solute.***
***24. How is the osmolality of a solution determined? What is a milliosmole?***

## Acids and Bases

Many molecules and compounds are classified as acids or bases. For most purposes an **acid** is defined as a proton donor. Because a hydrogen atom without its electron is a proton ($H^+$), any substance that releases hydrogen ions is an acid. Hydrochloric acid (HCl) forms hydrogen ions ($H^+$) and chloride ions ($Cl^-$) in solution and therefore is an acid.

$$HCl \rightarrow H^+ + Cl^-$$

A **base** is defined as a proton acceptor, and any substance that binds to (accepts) $H^+$ ions is a base. Many bases function as proton acceptors by releasing hydroxide ions ($OH^-$) when they dissociate. The base sodium hydroxide (NaOH) dissociates to form $Na^+$ and $OH^-$ ions.

$$NaOH \rightarrow Na^+ + OH^-$$

The $OH^-$ ions are proton acceptors that combine with $H^+$ ions to form water.

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

Acids and bases are classified as strong or weak. Strong acids or bases dissociate almost completely when dissolved in water. Consequently, they release almost all of their hydrogen or hydroxide ions. The more completely the acid or base dissociates, the stronger it is. For example, HCl is a strong acid because it completely dissociates in water.

$$HCl \rightarrow H^+ + Cl^-$$
Not freely reversible

Weak acids or bases only partially dissociate in water. Consequently, they release only some of their $H^+$ or $OH^-$ ions. For example, when acetic acid ($CH_3COOH$) is dissolved in water, some of it dissociates, but some of it remains in the undissociated form. An equilibrium is established between the ions and the undissociated weak acid.

$$CH_3COOH \rightleftarrows CH_3COO^- + H^+$$
Freely reversible

For a given weak acid or base, the amount of the dissociated ions relative to the weak acid or base is a constant.

### The pH Scale

The **pH scale** is a means of referring to the hydrogen ion concentration in a solution (figure 2.13). Pure water is defined as a **neutral solution** and has a pH of 7. A neutral solution has equal concentrations of hydrogen and hydroxide ions. Solutions with a pH less than 7 are **acidic** and have a greater concentration of hydrogen ions than hydroxide ions. **Alkaline** (al′kă-līn), or **basic,**

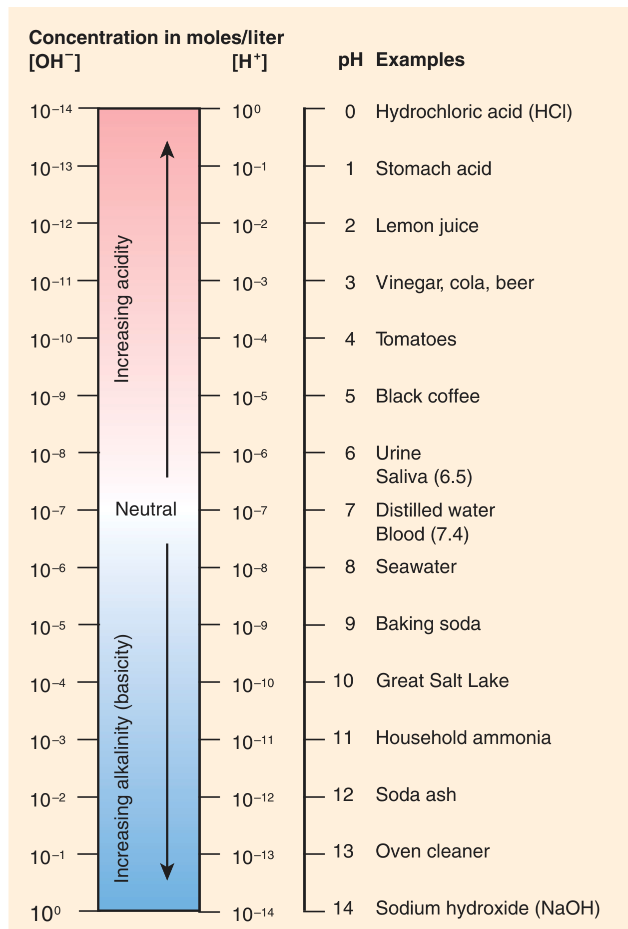



**Figure 2.13 The pH Scale**

A pH of 7 is considered neutral. Values less than 7 are acidic (the lower the number, the more acidic). Values greater than 7 are basic (the higher the number, the more basic). Representative fluids and their approximate pH values are listed.

solutions have a pH greater than 7 and have fewer hydrogen ions than hydroxide ions.

The symbol *pH* stands for power (p) of hydrogen ion ($H^+$) concentration. The power is a factor of 10, which means that a change in the pH of a solution by 1 pH unit represents a 10-fold change in the hydrogen ion concentration. For example, a solution of pH 6 has a hydrogen ion concentration 10 times greater than a solution of pH 7 and 100 times greater than a solution of pH 8. As the pH value becomes smaller, the solution has more hydrogen ions and is more acidic, and as the pH value becomes larger, the solution has fewer hydrogen ions and is more basic. Appendix D considers pH in greater detail.

### Acidosis and Alkalosis

The normal pH range for human blood is 7.35 to 7.45. **Acidosis** results if blood pH drops below 7.35, in which case the nervous system becomes depressed, and the individual can become disoriented and possibly comatose. **Alkalosis** results if blood pH rises above 7.45. Then the nervous system becomes overexcitable, and the individual can be extremely nervous or have convulsions. Both acidosis and alkalosis can be fatal.

### *Salts*

A **salt** is a compound consisting of a cation other than a hydrogen ion and an anion other than a hydroxide ion. Salts are formed by the interaction of an acid and a base in which the hydrogen ions of the acid are replaced by the positive ions of the base. For example, in a solution when hydrochloric acid (HCl) reacts with the base sodium hydroxide (NaOH), the salt, sodium chloride (NaCl), is formed.

$$\underset{\text{(Acid)}}{HCl} + \underset{\text{(Base)}}{NaOH} \rightarrow \underset{\text{(Salt)}}{NaCl} + \underset{\text{(Water)}}{H_2O}$$

Typically, when salts such as sodium chloride dissociate in water, they form positively and negatively charged ions (see figure 2.8).

### *Buffers*

The chemical behavior of many molecules changes as the pH of the solution in which they are dissolved changes. For example, many enzymes work best within narrow ranges of pH. The survival of an organism depends on its ability to regulate body fluid pH within a narrow range. Deviations from the normal pH range for human blood are life-threatening.

One way body fluid pH is regulated involves the action of buffers, which resist changes in solution pH when either acids or bases are added. A **buffer** is a solution of a conjugate acid–base pair in which the acid component and the base component occur in similar concentrations. A conjugate base is everything that remains of an acid after the hydrogen ion (proton) is lost. A conjugate acid is formed when a hydrogen ion is transferred to the conjugate base. Two substances related in this way are a **conjugate acid–base pair.** For example, carbonic acid ($H_2CO_3$) and bicarbonate ion ($HCO_3^-$), formed by the dissociation of carbonic acid, are a conjugate acid–base pair.

$$H_2CO_3 \rightleftarrows H^+ + HCO_3^-$$

In the forward reaction, carbonic acid loses a hydrogen ion to produce bicarbonate ion, which is a conjugate base. In the reverse reaction, a hydrogen ion is transferred to the bicarbonate ion (conjugate base) to produce carbonic acid, which is a conjugate acid.

For a given condition, this reversible reaction results in an equilibrium, in which the amounts of carbonic acid relative to the amounts of hydrogen ion and bicarbonate ions remains constant. The conjugate acid–base pair can resist changes in pH because of this equilibrium. If an acid is added to a buffer, the hydrogen ions from the added acid can combine with the base component of the conjugate acid–base pair. As a result, the concentration of hydrogen ions does not increase as much as it would without this reaction. If hydrogen ions are added to a carbonic acid solution, many of the hydrogen ions combine with bicarbonate ions to form carbonic acid.

On the other hand, if a base is added to a buffered solution, the conjugate acid can release hydrogen ions to counteract the effects of the added base. For example, if hydroxide ions are added to a carbonic acid solution, the hydroxide ions combine with hydrogen ions to form water. As the hydrogen ions are incorporated into water, carbonic acid dissociates to form hydrogen and bicarbonate ions, thereby maintaining the hydrogen ion concentration (pH) within a normal range.

The greater the buffer concentration, the more effective it is in resisting a change in pH, but buffers cannot entirely prevent some change in the pH of a solution. For example, when an acid is added to a buffered solution, the pH decreases but not to the extent it would have without the buffer. Several very important buffers are found in living systems and include bicarbonate, phosphates, amino acids, and proteins as components.

25. ***Define acid and base, and describe the pH scale. What is the difference between a strong acid or base and a weak acid or base?***
26. ***Define acidosis and alkalosis, and describe the symptoms of each.***
27. ***What is a salt? What is a buffer, and why are buffers important to organisms?***

**P R E D I C T 7**

**Dihydrogen phosphate ion ($H_2PO_4^-$) and monohydrogen phosphate ion ($HPO_4^{2-}$) form the phosphate buffer system.**

$$H_2PO_4^- \rightleftarrows H^+ + HPO_4^{2-}$$

**Identify the conjugate acid and conjugate base in the phosphate buffer system. Explain how they function as a buffer when either hydrogen or hydroxide ions are added to the solution.**

## Oxygen

**Oxygen** ($O_2$) is an inorganic molecule consisting of two oxygen atoms bound together by a double covalent bond. About 21% of the gas in the atmosphere is oxygen, and it is essential for most animals. Oxygen is required by humans in the final step of a series of reactions in which energy is extracted from food molecules (see chapters 3 and 25).

## Carbon Dioxide

**Carbon dioxide** ($CO_2$) consists of one carbon atom bound by double covalent bonds to two oxygen atoms. Carbon dioxide is produced when organic molecules such as glucose are metabolized within the cells of the body (see chapters 3 and 25). Much of the energy stored in the covalent bonds of glucose is transferred to other organic molecules when glucose is broken down, and carbon dioxide is released. Once carbon dioxide is produced, it is eliminated from the cell as a metabolic by-product, transferred to the lungs by blood, and exhaled during respiration. If carbon dioxide is allowed to accumulate within cells, it becomes toxic.

*28. What are the functions of oxygen and carbon dioxide in living systems?*

# Organic Chemistry

***Objectives***

- ***Describe the building blocks and functions of carbohydrates, lipids, proteins, and nucleic acids in the body.***
- ***Explain the function of ATP in the body.***

The ability of carbon to form covalent bonds with other atoms makes possible the formation of the large, diverse, complicated molecules necessary for life. A series of carbon atoms bound together by covalent bonds constitutes the "backbone" of many large molecules. Variation in the length of the carbon chains and the combination of atoms bound to the carbon backbone allows for the formation of a wide variety of molecules. For example, some protein molecules have thousands of carbon atoms bound by covalent bonds to one another or to other atoms, such as nitrogen, sulfur, hydrogen, and oxygen.

The four major groups of organic molecules essential to living organisms are carbohydrates, lipids, proteins, and nucleic acids. Each of these groups has specific structural and functional characteristics.

## Carbohydrates

**Carbohydrates** are composed primarily of carbon, hydrogen, and oxygen atoms and range in size from small to very large. In most carbohydrates, for each carbon atom there are approximately two hydrogen atoms and one oxygen atom. Note that the ratio of hydrogen atoms to oxygen atoms is two to one, the same as in water. They are called carbohydrates because carbon (carbo) atoms are combined with the same atoms that form a water molecule (hydrated). The large number of oxygen atoms in carbohydrates makes them relatively polar molecules. Consequently, they are soluble in polar solvents such as water.

Carbohydrates are important parts of other organic molecules, and they can be broken down to provide the energy necessary for life. Undigested carbohydrates also provide bulk in feces, which helps to maintain the normal function and health of the digestive tract. Table 2.5 summarizes the roles of carbohydrates in the body.

**Table 2.5 Role of Carbohydrates in the Body**

| Role | Example |
|---|---|
| Structure | Ribose forms part of RNA and ATP molecules, and deoxyribose forms part of DNA. |
| Energy | Monosaccharides (glucose, fructose, galactose) can be used as energy sources. Disaccharides (sucrose, lactose, maltose) and polysaccharides (starch, glycogen) must be broken down to monosaccharides before they can be used for energy. Glycogen is an important energy-storage molecule in muscles and in the liver. |
| Bulk | Cellulose forms bulk in the feces. |

### *Monosaccharides*

Large carbohydrates are composed of numerous, relatively simple building blocks called **monosaccharides** (mon-ō-sak′ă-rīdz; the prefix *mono-* means one; the term *saccharide* means sugar), or simple sugars. Monosaccharides commonly contain three carbons (trioses), four carbons (tetroses), five carbons (pentoses), or six carbons (hexoses).

The monosaccharides most important to humans include both five- and six-carbon sugars. Common six-carbon sugars, such as glucose, fructose, and galactose, are **isomers** (ī′sō-merz), which are molecules that have the same number and types of atoms but differ in their three-dimensional arrangement (figure 2.14). Glucose, or blood sugar, is the major carbohydrate found in the blood and is a major nutrient for most cells of the body. Fructose and galactose are also important dietary nutrients. Important five-carbon sugars include ribose and deoxyribose (see figure 2.24), which are components of ribonucleic acid (RNA) and deoxyribonucleic acid (DNA), respectively.

### *Disaccharides*

**Disaccharides** (dī-sak′ă-rīdz; *di-* means two) are composed of two simple sugars bound together through a dehydration reaction. Glucose and fructose, for example, combine to form a disaccharide called **sucrose** (table sugar) plus a molecule of water (figure 2.15*a*). Several disaccharides are important to humans, including sucrose, lactose, and maltose. Lactose, or milk sugar, is glucose combined with galactose; and maltose, or malt sugar, is two glucose molecules joined together.

### *Polysaccharides*

**Polysaccharides** (pol-ē-sak′ă-rīdz; *poly-* means many) consist of many monosaccharides bound together to form long chains that are either straight or branched. **Glycogen,** or animal starch, is a polysaccharide composed of many glucose molecules (figure 2.15*b*). Because glucose can be metabolized rapidly and the resulting energy can be used by cells, glycogen is an important energy-storage molecule. A substantial amount of the glucose that is metabolized to produce energy for muscle contraction during exercise is stored in the form of glycogen in the cells of the liver and skeletal muscles.

**Starch** and **cellulose** are two important polysaccharides found in plants, and both are composed of long chains of glucose. Plants use starch as an energy storage molecule in the same way that animals use glycogen, and cellulose is an important structural component of plant cell walls. When humans ingest plants, the starch can be broken down and used as an energy source. Humans, however, do not have the digestive enzymes necessary to break down cellulose. The cellulose is eliminated in the feces, where it provides bulk.

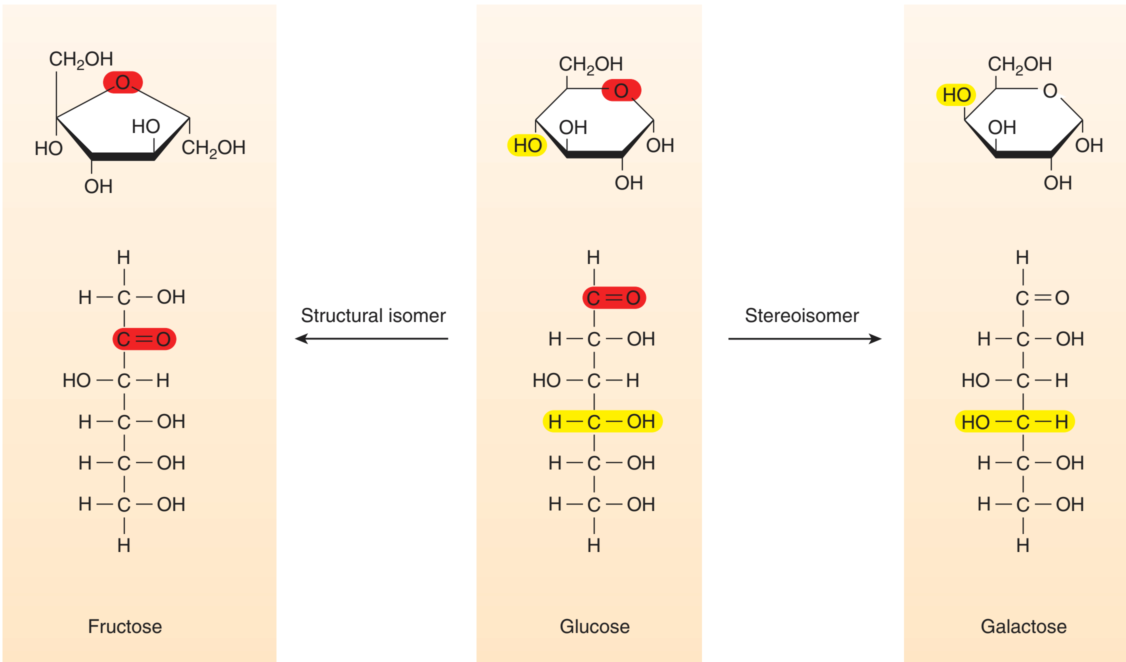



**Figure 2.14** **Monosaccharides**

These monosaccharides almost always form a ring-shaped molecule. They are represented as linear models to more readily illustrate the relationships between the atoms of the molecules. Fructose is a structural isomer of glucose because it has identical chemical groups bonded in a different arrangement in the molecule (*indicated by red shading*). Galactose is a stereoisomer of glucose because it has exactly the same groups bonded to each carbon atom but located in a different three-dimensional orientation (*indicated by yellow shading*).

## Lipids

**Lipids** are a second major group of organic molecules common to living systems. Like carbohydrates, they are composed principally of carbon, hydrogen, and oxygen; but other elements, such as phosphorus and nitrogen, are minor components of some lipids. Lipids contain a lower ratio of oxygen to carbon than do carbohydrates, which makes them less polar. Consequently, lipids can be dissolved in nonpolar organic solvents, such as alcohol or acetone, but they are relatively insoluble in water.

Lipids have many important functions in the body. They provide protection and insulation, help to regulate many physiologic processes, and form plasma membranes. In addition, lipids are a major energy storage molecule and can be broken down and used as a source of energy. Table 2.6 summarizes the many roles of lipids in the body. Several different kinds of molecules, such as fats, phospholipids, steroids, and prostaglandins, are classified as lipids.

**Fats** are a major type of lipid. Like carbohydrates, fats are ingested and broken down by hydrolysis reactions in cells to release energy for use by those cells. Conversely, if intake exceeds need, excess chemical energy from any source can be stored in the body as fat for later use as energy is needed. Fats also provide protection by surrounding and padding organs, and under-the-skin fats act as an insulator to prevent heat loss.

**Table 2.6 Role of Lipids in the Body**

| Role | Example |
|---|---|
| Protection | Fat surrounds and pads organs. |
| Insulation | Fat under the skin prevents heat loss. Myelin surrounds nerve cells and electrically insulates the cells from one another. |
| Regulation | Steroid hormones regulate many physiologic processes. For example, estrogen and testosterone are sex hormones responsible for many of the differences between males and females. Prostaglandins help regulate tissue inflammation and repair. |
| Vitamins | Fat-soluble vitamins perform a variety of functions. Vitamin A forms retinol, which is necessary for seeing in the dark; active vitamin D promotes calcium uptake by the small intestine; vitamin E promotes wound healing; and vitamin K is necessary for the synthesis of proteins responsible for blood clotting. |
| Structure | Phospholipids and cholesterol are important components of plasma membranes. |
| Energy | Lipids can be stored and broken down later for energy; per unit of weight, they yield more energy than carbohydrates or proteins. |

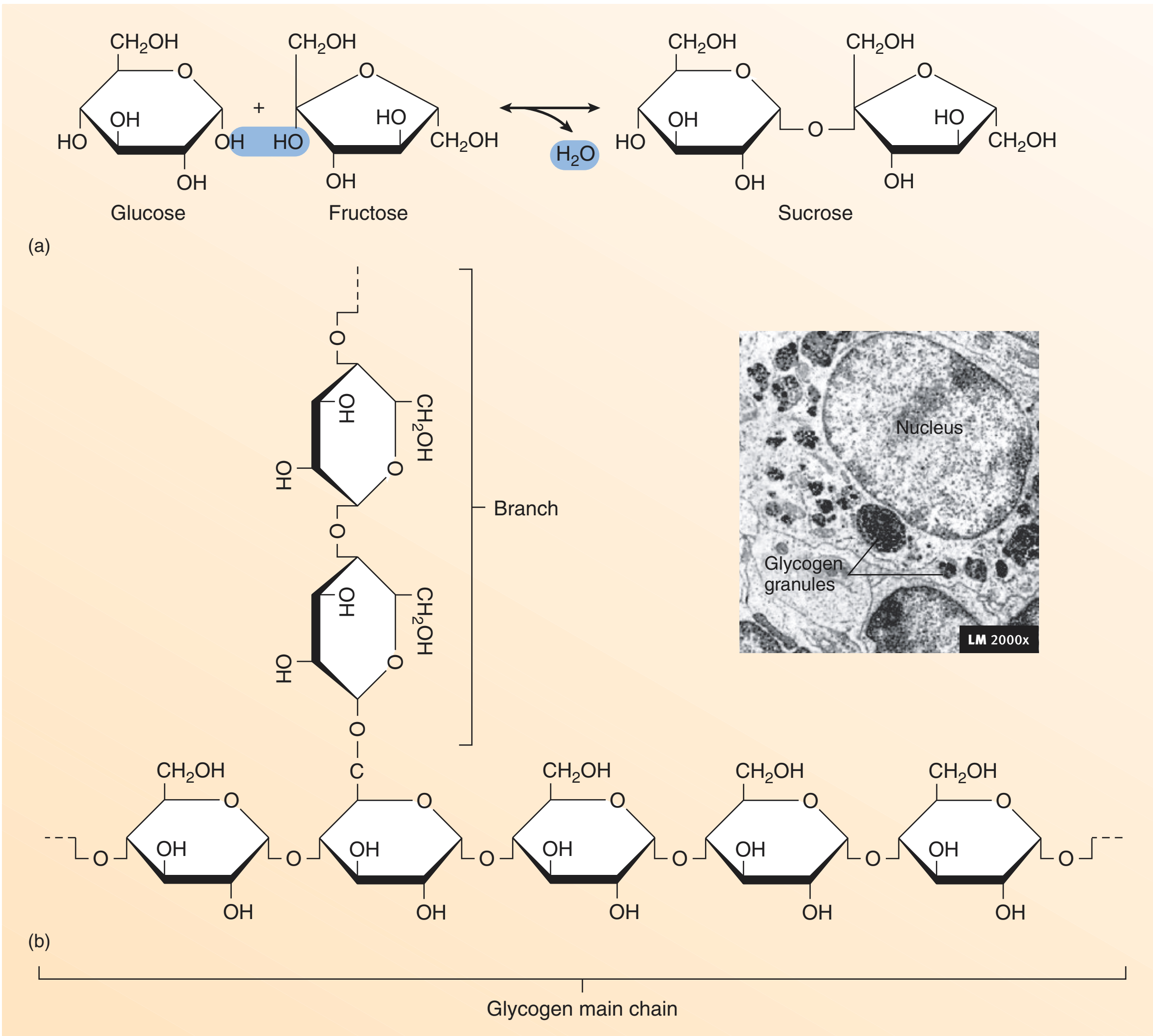



**Figure 2.15 Disaccharide and Polysaccharide**

(*a*) Formation of sucrose, a disaccharide, by a dehydration reaction involving glucose and fructose (monosaccharides). (*b*) Glycogen is a polysaccharide formed by combining many glucose molecules. The photo shows glycogen granules in a liver cell.

**Triglycerides** (trī-glis′er-īdz) constitute 95% of the fats in the human body. Triglycerides, which are sometimes called **triacylglycerols** (tri-as′il-glis′er-olz), consist of two different types of building blocks: one glycerol and three fatty acids. **Glycerol** is a three-carbon molecule with a hydroxyl group attached to each carbon atom, and **fatty acids** consist of a straight chain of carbon atoms with a carboxyl group attached at one end (figure 2.16). A **carboxyl** (kar-bok′sil) **group** (—COOH) consists of both an oxygen atom and a hydroxyl group attached to a carbon atom. The carboxyl group is responsible for the acidic nature of the molecule because it releases hydrogen ions into solution. Glycerides can be described according to the number and kinds of fatty acids that combine with glycerol through dehydration reactions. Monoglycerides have one fatty acid, diglycerides have two fatty acids, and triglycerides have three fatty acids bound to glycerol.

Fatty acids differ from one another according to the length and the degree of saturation of their carbon chains. Most naturally occurring fatty acids contain an even number of carbon atoms, with 14- to 18-carbon chains being the most common. A fatty acid is **saturated** (figure 2.17) if it contains only single covalent bonds between the carbon atoms. Sources of saturated fats include beef, pork, whole milk, cheese, butter, eggs, coconut oil, and palm oil. The carbon chain is **unsaturated** if it has one or more double covalent bonds between carbon atoms. Because the double covalent bonds

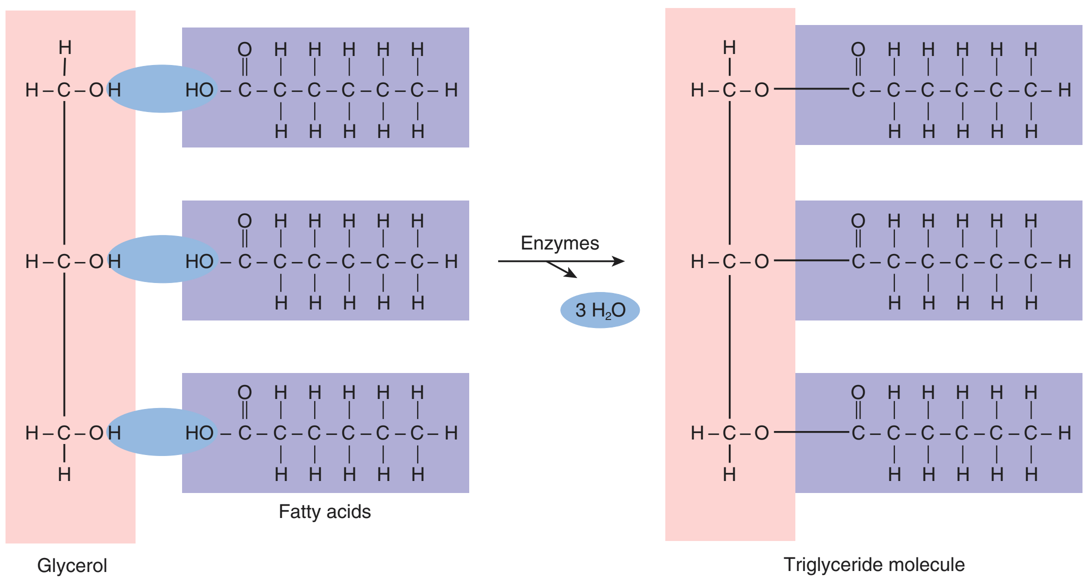



**Figure 2.16 Triglyceride**
Production of a triglyceride from one glycerol molecule and three fatty acids.

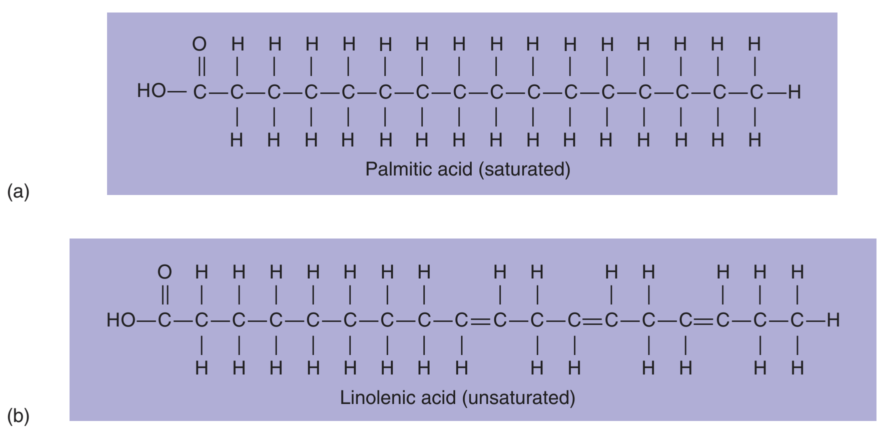
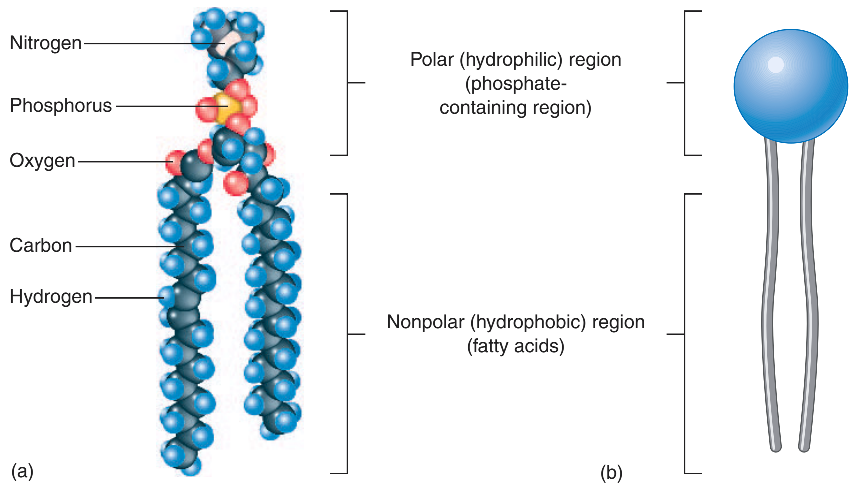



**Figure 2.17 Fatty Acids**
(*a*) Palmitic acid (saturated with no double bonds between the carbons). (*b*) Linolenic acid (unsaturated with three double bonds between the carbons).

can occur anywhere along the carbon chain, many types of unsaturated fatty acids with an equal degree of unsaturation are possible. **Monounsaturated** fats, such as olive and peanut oils, have one double covalent bond between carbon atoms. **Polyunsaturated fats,** such as safflower, sunflower, corn, or fish oils, have two or more double covalent bonds between carbon atoms. Unsaturated fats are the best type of fats in the diet because unlike saturated fats they do not contribute to the development of cardiovascular disease.

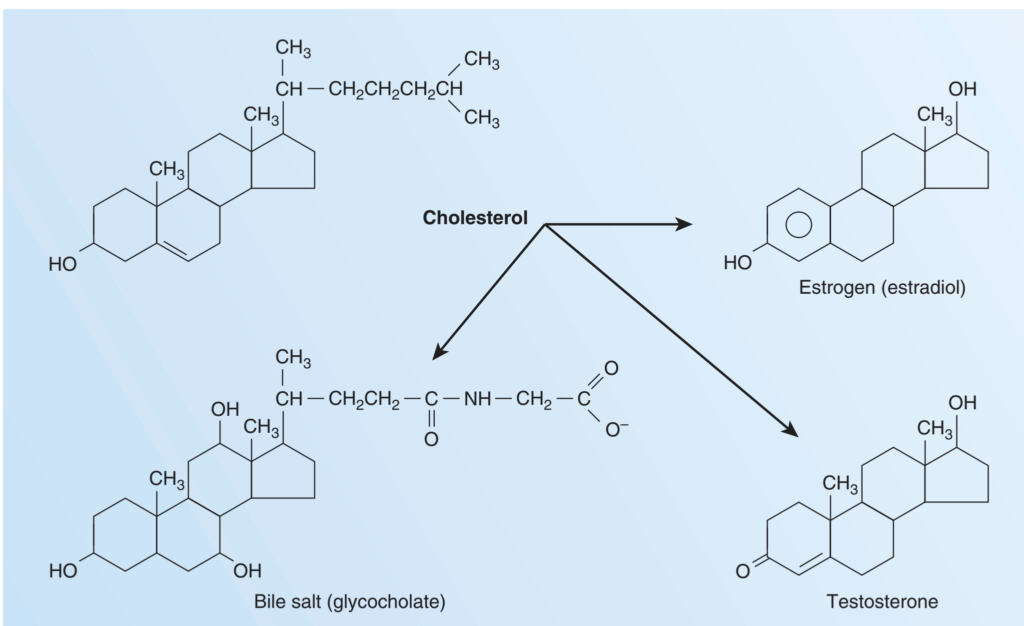
**Phospholipids** are similar to triglycerides, except that one of the fatty acids bound to the glycerol is replaced by a molecule containing phosphate and, usually, nitrogen (figure 2.18). They are polar at the end of the molecule to which the phosphate is bound and nonpolar at the other end. The polar end of the molecule is attracted to water, and the nonpolar end is repelled by water. Phospholipids are important structural components of plasma membranes (see chapter 3).

The **eicosanoids** (ī′kō-să-noydz) are a group of important chemicals derived from fatty acids. They include **prostaglandins** (pros′tă-glan′dinz), **thromboxanes** (throm′bok-zānz), and **leukotrienes** (loo-kō-trī′ēnz). Eicosanoids are made in most cells and are important regulatory molecules. Among their numerous effects is their role in the response of tissues to injuries. Prostaglandins have been implicated in regulating the secretion of some hormones, blood clotting, some reproductive functions, and many other processes. Many of the therapeutic effects of aspirin and other anti-inflammatory drugs result from their ability to inhibit prostaglandin synthesis.




## Figure 2.18 Phospholipids

(*a*) Molecular model of a phospholipid. (*b*) Simplified way in which phospholipids are often depicted.




## Figure 2.19 Steroids

Steroids are four-ringed molecules that differ from one another according to the groups attached to the rings. Cholesterol, the most common steroid, can be modified to produce other steroids.

**Steroids** differ in chemical structure from other lipid molecules, but their solubility characteristics are similar. All steroid molecules are composed of carbon atoms bound together into four ringlike structures (figure 2.19). Important steroid molecules include cholesterol, bile salts, estrogen, progesterone, and testosterone. Cholesterol is an important steroid because other molecules are synthesized from it. For example, bile salts, which increase fat absorption in the intestines, are derived from cholesterol, as are the reproductive hormones estrogen, progesterone, and testosterone. In addition, cholesterol is an important component of plasma membranes. Although high levels of cholesterol in the blood increase the risk of cardiovascular disease, a certain amount of cholesterol is vital for normal function.

Another class of lipids is the **fat-soluble vitamins.** Their structures are not closely related to one another, but they are nonpolar molecules essential for many normal functions of the body.

## Proteins

All **proteins** contain carbon, hydrogen, oxygen, and nitrogen bound together by covalent bonds, and most proteins contain some sulfur. In addition, some proteins contain small amounts of phosphorus, iron, and iodine. The molecular mass of proteins can be very large. For the purpose of comparison, the molecular mass of water is approximately 18, sodium chloride 58, and glucose 180; but the molecular mass of proteins ranges from approximately 1000 to several million.

Proteins regulate bodily processes, act as a transportation system in the body, provide protection, help muscles contract, and provide structure and energy. Table 2.7 summarizes the functions of proteins in the body.

### *Protein Structure*

The basic building blocks for proteins are the 20 **amino** (ă-mē′nō) **acid** molecules. Each amino acid has an amine (ă-mēn′) group ($—NH_2$), a carboxyl group ($—COOH$), a hydrogen atom, and a side chain designated by the symbol *R* attached to the same carbon atom. The side chain can be a variety of chemical structures, and the differences in the side chains make the amino acids different from one another (figure 2.20).

Covalent bonds formed between amino acid molecules during protein synthesis are called **peptide bonds** (figure 2.21). A dipeptide is two amino acids bound together by a peptide bond, a tripeptide is three amino acids bound together by peptide bonds, and a polypeptide is many amino acids bound together by peptide bonds. Proteins are polypeptides composed of hundreds of amino acids.

The **primary structure** of a protein is determined by the sequence of the amino acids bound by peptide bonds (figure 2.22*a*).

**Table 2.7 Role of Proteins in the Body**

| Role | Example |
|---|---|
| Regulation | Enzymes control chemical reactions. Hormones regulate many physiologic processes; for example, insulin affects glucose transport into cells. |
| Transport | Hemoglobin transports oxygen and carbon dioxide in the blood. Plasma proteins transport many substances in the blood. Proteins in plasma membranes control the movement of materials into and out of the cell. |
| Protection | Antibodies and complement protect against microorganisms and other foreign substances. |
| Contraction | Actin and myosin in muscle are responsible for muscle contraction. |
| Structure | Collagen fibers form a structural framework in many parts of the body. Keratin adds strength to skin, hair, and nails. |
| Energy | Proteins can be broken down for energy; per unit of weight, they yield as much energy as carbohydrates. |

The potential number of different protein molecules is enormous because 20 different amino acids exist and each amino acid can be located at any position along a polypeptide chain. The characteristics of the amino acids in a protein ultimately determine the three-dimensional shape of the protein, and the shape of the protein determines its function. A change in one, or a few, amino acids in the primary structure can alter protein function, usually making the protein less or even nonfunctional.

The **secondary structure** results from the folding or bending of the polypeptide chain caused by the hydrogen bonds between amino acids (figure 2.22*b*). Two common shapes that result are helices or pleated sheets. If the hydrogen bonds that maintain the shape of the protein are broken, the protein becomes nonfunctional. This change in shape is called **denaturation,** and it can be

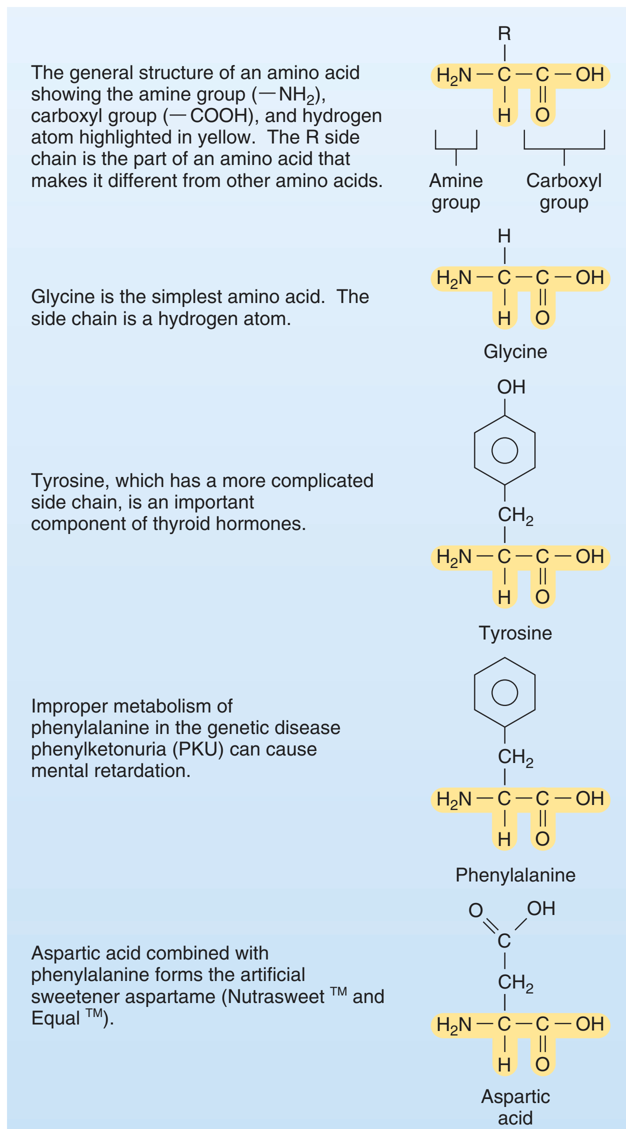



**Figure 2.20** Amino Acids

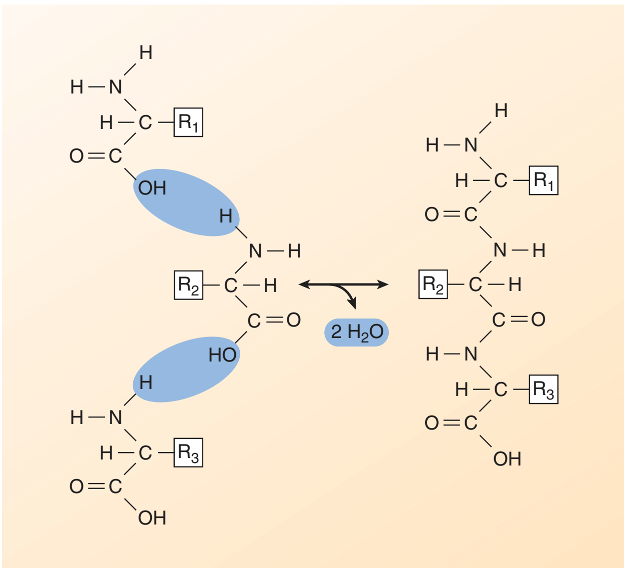



**Figure 2.21** **Peptide Bonds**

Dehydration reaction between three amino acids (*left*) to form a tripeptide (*right*). One water molecule ($H_2O$) is given off for each peptide bond formed.

caused by abnormally high temperatures or changes in the pH of body fluids. An everyday example of denaturation is the change in the proteins of egg whites when they are cooked.

The **tertiary structure** results from the folding of the helices or pleated sheets (figure 2.22*c*). Some amino acids are quite polar and therefore form hydrogen bonds with water. The polar portions of proteins tend to remain unfolded, maximizing their contact with water, whereas the less polar regions tend to fold into a globular shape, minimizing their contact with water. The formation of covalent bonds between sulfur atoms of one amino acid and sulfur atoms in another amino acid located at a different place in the sequence of amino acids can also contribute to the tertiary structure of proteins. The tertiary structure determines the shape of a **domain,** which is a folded sequence of 100–200 amino acids within a protein. The functions of proteins occur at one or more domains. Therefore, changes in the primary or secondary structure that affect the shape of the domain can change protein function.

If two or more proteins associate to form a functional unit, the individual proteins are called subunits. The **quaternary structure** refers to the spatial relationships between the individual subunits (figure 2.22*d*).

### *Enzymes*

Proteins perform many roles in the body, including acting as enzymes. An **enzyme** is a protein catalyst that increases the rate at which a chemical reaction proceeds without the enzyme being permanently changed. The three-dimensional shape of enzymes is critical for their normal function because it determines the structure of the enzyme's **active site.** According to the **lock-and-key model** of enzyme action, a reaction occurs when the reactants (key) bind to the active site (lock) on the enzyme. This view of enzymes and reactants as rigid structures fitting together has been modified by the **induced fit model,** in which the enzyme is able to slightly change shape and better fit the reactants. The enzyme is like a glove that does not achieve its functional shape until the hand (reactants) moves into place.

At the active site, reactants are brought into close proximity (figure 2.23). After the reactants combine, they are released from the active site, and the enzyme is capable of catalyzing additional reactions. The activation energy required for a chemical reaction to occur is lowered by enzymes (see figure 2.12) because they orient the reactants toward each other in such a way that it is more likely a chemical reaction will occur.

Slight changes in the structure of an enzyme can destroy the ability of the active site to function. Enzymes are very sensitive to changes in temperature or pH, which can break the hydrogen bonds within them. As a result, the relationship between amino acids changes, thereby producing a change in shape that prevents the enzyme from functioning normally.

To be functional, some enzymes require additional, nonprotein substances called **cofactors.** The cofactor can be an ion, such as magnesium or zinc, or an organic molecule. Cofactors that are organic molecules, such as certain vitamins, may be referred to as **coenzymes.** Cofactors normally form part of the enzyme's active site and are required to make the enzyme functional.

Enzymes are highly specific because their active site can bind only to certain reactants. Each enzyme catalyzes a specific chemical reaction and no others. Many different enzymes are therefore needed to catalyze the many chemical reactions of the body. Enzymes often are named by adding the suffix *-ase* to the name of the molecules on which they act. For example, an enzyme that catalyzes the breakdown of lipids is a **lipase** (lip′ās, lī′pās), and an enzyme that breaks down proteins is called a **protease** (prō′tē-ās).

Enzymes control the rate at which most chemical reactions proceed in living systems. Consequently, they control essentially all cellular activities. At the same time, the activity of enzymes themselves is regulated by several mechanisms that exist within the cells. Some mechanisms control the enzyme concentration by influencing the rate at which the enzymes are synthesized, and others alter the activity of existing enzymes. Much of what is known about the regulation of cellular activity involves knowledge of how enzyme activity is controlled.

## Nucleic Acids: DNA and RNA

**Deoxyribonucleic** (dē-oks′ē-rī′bō-noo-klē′ik) **acid (DNA)** is the genetic material of cells, and copies of DNA are transferred from one generation of cells to the next generation. DNA contains the information that determines the structure of proteins. **Ribonucleic** (rī′bō-noo-klē′ik) **acid (RNA)** is structurally related to DNA, and three types of RNA also play important roles in protein synthesis. In chapter 3 the means by which DNA and RNA direct the functions of the cell are described.

The **nucleic** (noo-klē′ik, noo-klā′ik) **acids** are large molecules composed of carbon, hydrogen, oxygen, nitrogen, and phosphorus. Both DNA and RNA consist of basic building blocks called **nucleotides** (noo′klē-ō-tīdz). Each nucleotide is composed of a monosaccharide to which a nitrogenous organic base and a

(a) Primary structure—the amino acid sequence

(b) Secondary structure with folding as a result of hydrogen bonding (*dotted red lines*)

(c) Tertiary structure with secondary folding caused by interactions within the polypeptide and its immediate environment

(d) Quaternary structure—the relationships between individual subunits

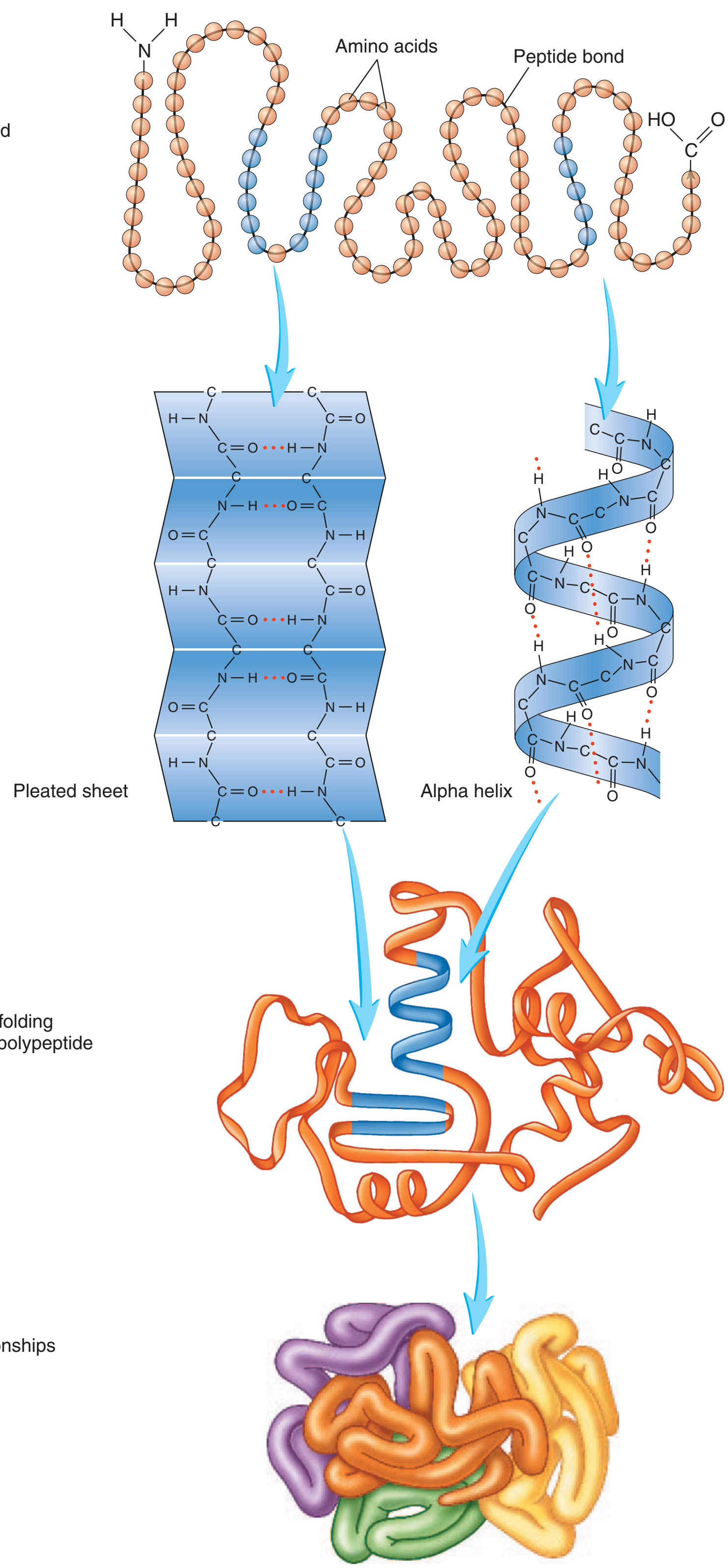
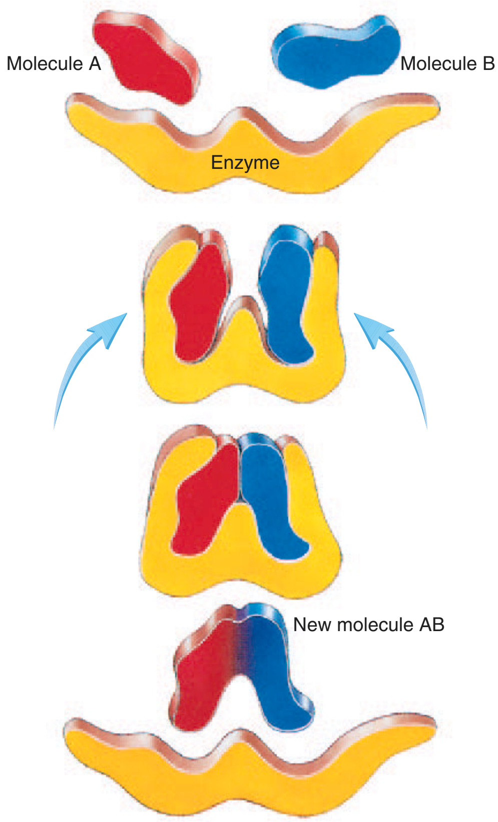
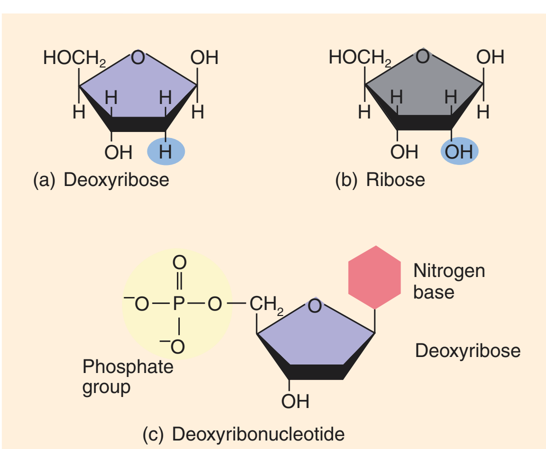



**Figure 2.22** Protein Structure




## Figure 2.23 Enzyme Action

The enzyme brings the two reacting molecules together. This is possible because the reacting molecules "fit" the shape of the enzyme (lock-and-key model). After the reaction, the unaltered enzyme can be used again.



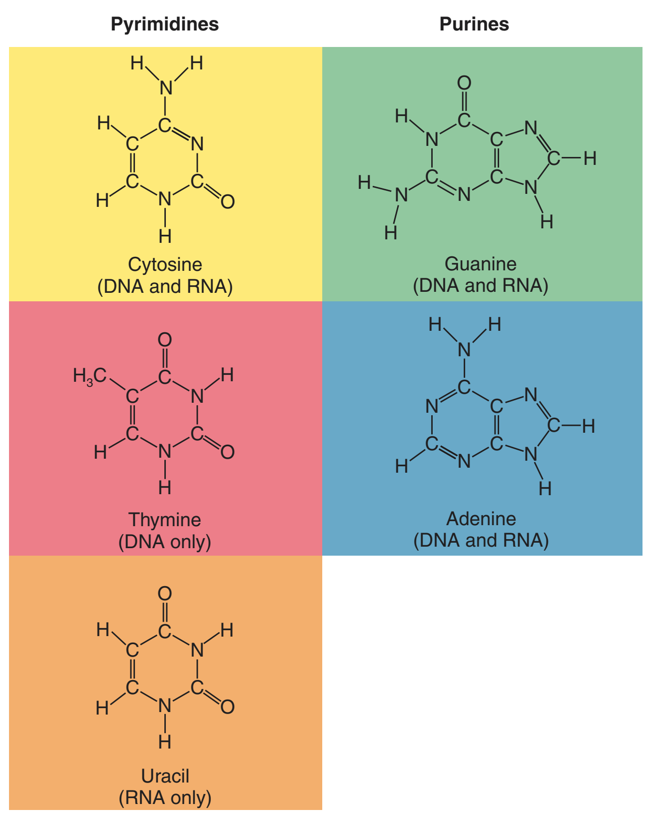

## Figure 2.24 Components of Nucleotides

(*a*) Deoxyribose sugar, which forms nucleotides used in DNA production. (*b*) Ribose sugar, which forms nucleotides used in RNA production. Note that deoxyribose is ribose minus an oxygen atom. (*c*) Deoxyribonucleotide consisting of deoxyribose, a nitrogen base, and a phosphate group.

phosphate group are attached (figure 2.24). The monosaccharide is deoxyribose for DNA, and ribose for RNA. The organic bases are thymine (thī′mēn, thī′min), cytosine (sī′tō-sēn), and uracil (ūr′ă-sil), which are single-ringed pyrimidines (pī-rim′i-dēnz); and adenine (ad′ĕ-nēn) and guanine (gwahn′ēn), which are double-ringed purines (pūr′ēnz) (figure 2.25).

DNA has two strands of nucleotides joined together to form a twisted ladderlike structure called a double helix (figure 2.26). The uprights of the ladder are formed by covalent bonds between the deoxyribose molecules and phosphate groups of adjacent nucleotides. The rungs of the ladder are formed by the bases of the nucleotides of one upright connected to the bases of the other upright by hydrogen bonds. Each nucleotide of DNA contains one of the organic bases: adenine, thymine, cytosine, or guanine. Adenine binds only to thymine because the structure of these organic bases allows two hydrogen bonds to form between them. Cytosine binds only to guanine because the structure of these organic bases allows three hydrogen bonds to form between them.

The sequence of organic bases in DNA molecules stores genetic information. Each DNA molecule consists of millions of organic bases, and their sequence ultimately determines the type and sequence of amino acids found in protein molecules. Because enzymes are proteins, DNA structure determines the rate and type




## Figure 2.25 Nitrogenous Organic Bases

The organic bases found in nucleic acids are separated into two groups. Purines are double-ringed molecules, and pyrimidines are single-ringed molecules.

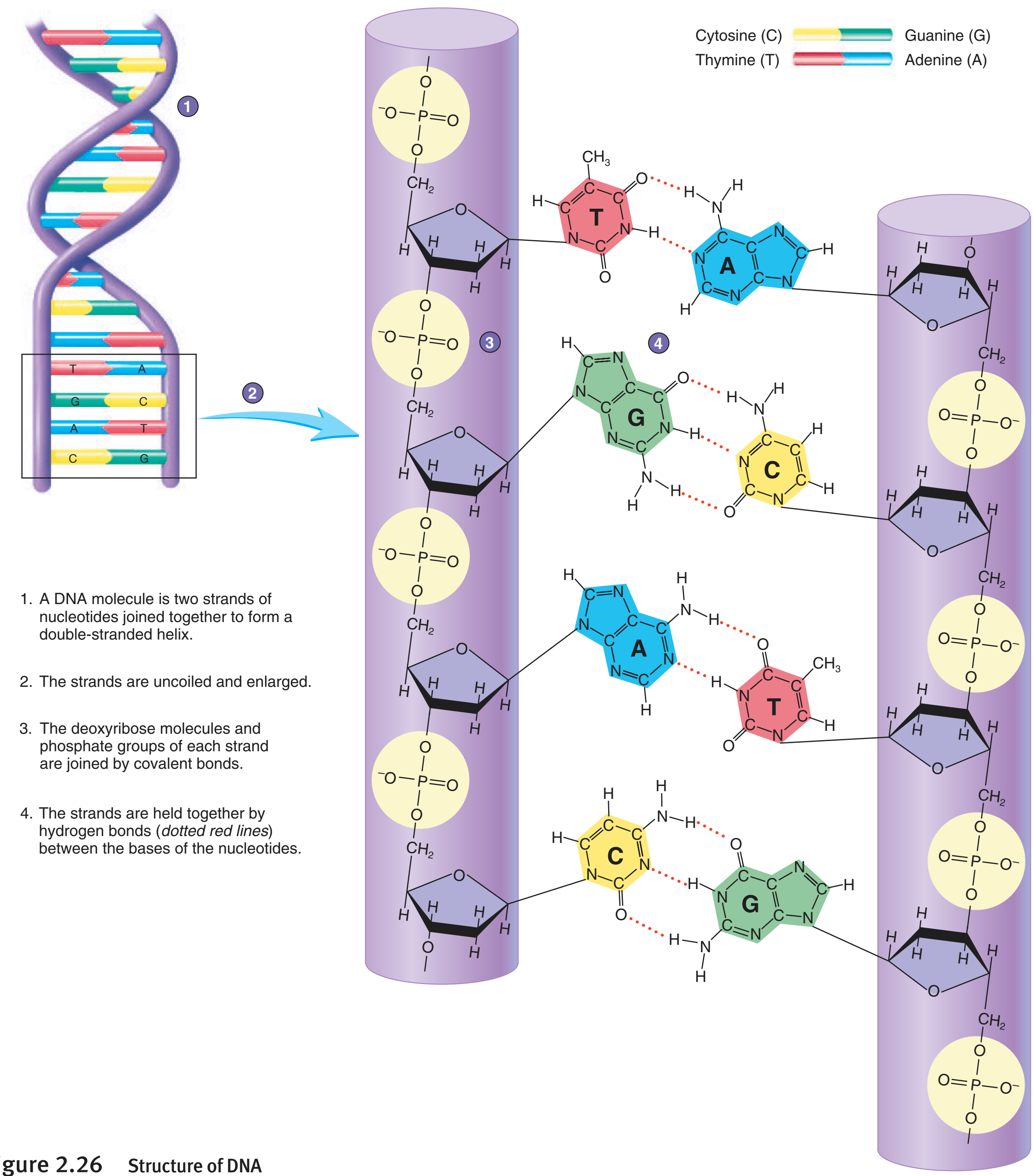



**Figure 2.26** **Structure of DNA**

of chemical reactions that occur in cells by controlling enzyme structure. The information contained in DNA therefore ultimately defines all cellular activities. Other proteins, such as collagen, that are coded by DNA determine many of the structural features of humans.

RNA has a structure similar to a single strand of DNA. Like DNA, four different nucleotides make up the RNA molecule, and the organic bases are the same, except that thymine is replaced with uracil (see figure 2.25). Uracil can bind only to adenine.

## Adenosine Triphosphate

**Adenosine triphosphate** (ă-den′ō-sēn trī-fos′fāt) **(ATP)** is an especially important organic molecule found in all living organisms. It consists of adenosine and three phosphate groups (figure 2.27). Adenosine is the sugar ribose with the organic base adenine. The potential energy stored in the covalent bond between the second and third phosphate groups is important to living organisms because it provides the energy used in nearly all of the chemical reactions within cells.

The catabolism of glucose and other nutrient molecules results in chemical reactions that release energy. Some of that energy is used to synthesize ATP from ADP and an inorganic phosphate group ($P_i$):

$$\text{ADP} + P_i + \text{Energy (from catabolism)} \rightarrow \text{ATP}$$

The transfer of energy from nutrient molecules to ATP involves a series of oxidation–reduction reactions in which a high-energy electron is transferred from one molecule to the next molecule in the series. In chapter 25 the oxidation–reduction reactions of metabolism are considered in greater detail.

Once produced, ATP is used to provide energy for other chemical reactions (anabolism) or to drive cell processes such as muscle contraction. In the process, ATP is converted back to ADP and an inorganic phosphate group.

$$\text{ATP} \rightarrow \text{ADP} + P_i + \text{Energy (for anabolism and other cell processes)}$$

ATP is often called the energy currency of cells because it is capable of both storing and providing energy. The concentration of ATP is maintained within a narrow range of values, and essentially all energy-requiring chemical reactions stop when there is an inadequate quantity of ATP.

*29. List the four types of organic molecules important to life.*
*30. Name the basic building blocks of carbohydrates, fats, proteins, and nucleic acids.*
*31. List three types of carbohydrates, and explain the role of each in the body.*
*32. Distinguish between fats, phospholipids, and steroids, and give an example of each. What is a saturated fat?*
*33. Define a peptide bond. What makes proteins different from one another?*
*34. What determines the primary, secondary, tertiary, and quaternary structures of proteins? Define denaturation and name two things that can cause it to occur.*
*35. Compare the lock-and-key model and the induced fit model of enzyme activity. Define cofactor and coenzyme.*
*36. What are the structural and functional differences between DNA and RNA?*
*37. Describe the structure of ATP. What role does this molecule play in energy exchange?*

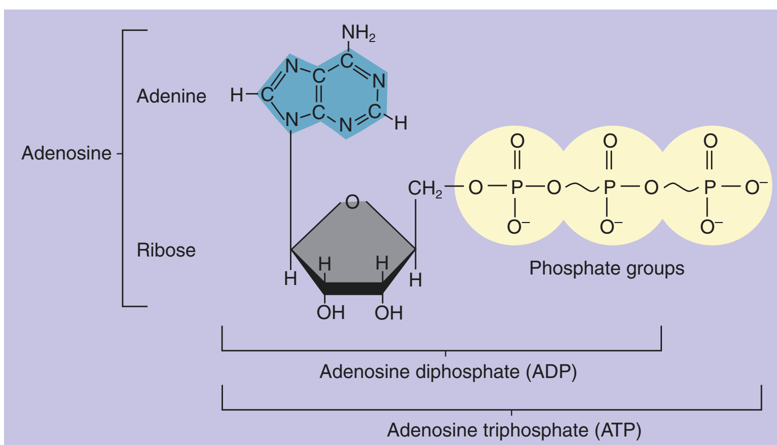



**Figure 2.27** Adenosine Triphosphate (ATP) Molecule

## S U M M A R Y

Chemistry is the study of the composition, structure, and properties of substances and the reactions they undergo. Much of the structure and function of healthy or diseased organisms can be understood at the chemical level.

### Basic Chemistry (p. 27)

#### Matter, Mass, and Weight

1. Matter is anything that occupies space.
2. Mass is the amount of matter in an object.
3. Weight results from the force exerted by earth's gravity on matter.

#### Elements and Atoms

1. An element is the simplest type of matter with unique chemical and physical properties.
2. An atom is the smallest particle of an element that has the chemical characteristics of that element. An element is composed of only one kind of atom.

3. Atoms consist of protons, neutrons, and electrons.
   - Protons are positively charged, electrons are negatively charged, and neutrons have no charge.
   - Protons and neutrons are found in the nucleus, and electrons, which are located around the nucleus, can be represented by an electron cloud.
4. The atomic number is the unique number of protons in an atom. The mass number is the sum of the protons and the neutrons.
5. Isotopes are atoms that have the same atomic number but different mass numbers.
6. The atomic mass of an element is the average mass of its naturally occurring isotopes weighted according to their abundance.

### Electrons and Chemical Bonding

1. The chemical behavior of atoms is determined mainly by their outermost electrons. A chemical bond occurs when atoms share or transfer electrons.
2. Ions are atoms that have gained or lost electrons.
   - An atom that loses one or more electrons becomes positively charged and is called a cation. An anion is an atom that becomes negatively charged after accepting one or more electrons.
   - Ionic bonding is the attraction of the oppositely charged cation and anion to each other.
3. A covalent bond is the sharing of electron pairs between atoms. A polar covalent bond results when the sharing of electrons is unequal and can produce a polar molecule that is electrically asymmetric.

### Molecules and Compounds

1. A molecule is two or more atoms chemically combined to form a structure that behaves as an independent unit. A compound is two or more *different* types of atoms chemically combined.
2. The kinds and numbers of atoms (or ions) in a molecule or compound can be represented by a formula consisting of the symbols of the atoms (or ions) plus subscripts denoting the number of each type of atom (or ion).
3. The molecular mass of a molecule or compound can be determined by adding up the atomic masses of its atoms (or ions).

### Intermolecular Forces

1. A hydrogen bond is the weak attraction that occurs between the oppositely charged regions of polar molecules. Hydrogen bonds are important in determining the three-dimensional structure of large molecules.
2. Solubility is the ability of one substance to dissolve in another. Ionic substances that dissolve in water by dissociation are electrolytes. Molecules that do not dissociate are nonelectrolytes.

## Chemical Reactions and Energy (p. 34)

### Synthesis Reactions

1. Synthesis reactions are the chemical combination of two or more substances to form a new or larger substance.
2. Dehydration reactions are synthesis reactions in which water is produced.
3. Anabolism is the sum of all the synthesis reactions in the body.

### Decomposition Reactions

1. Decomposition reactions are the chemical breakdown of a larger substance to two or more different smaller substances.
2. Hydrolysis reactions are decomposition reactions in which water is depleted.
3. All of the decomposition reactions in the body are called catabolism.

### Reversible Reactions

Reversible reactions produce an equilibrium condition in which the amount of reactants relative to the amount of products remains constant.

### Oxidation–Reduction Reactions

Oxidation–reduction reactions involve the complete or partial transfer of electrons between atoms.

### Energy

Energy is the ability to do work. Potential energy is stored energy, and kinetic energy is energy resulting from movement of an object.

### Chemical Energy

1. Chemical bonds are a form of potential energy.
2. Chemical reactions in which the products contain more potential energy than the reactants require the input of energy.
3. Chemical reactions in which the products have less potential energy than the reactants release energy.

### Heat Energy

1. Heat energy is energy that flows between objects that are at different temperatures.
2. Heat energy is released in chemical reactions and is responsible for body temperature.

### Speed of Chemical Reactions

1. Activation energy is the minimum energy that the reactants must have to start a chemical reaction.
2. Enzymes are specialized protein catalysts that lower the activation energy for chemical reactions. Enzymes speed up chemical reactions but are not consumed or altered in the process.
3. Increased temperature and concentration of reactants can increase the rate of chemical reactions.

## Inorganic Chemistry (p. 39)

Inorganic chemistry is mostly concerned with noncarbon-containing substances but does include some carbon-containing substances, such as carbon dioxide and carbon monoxide.

### Water

1. Water is a polar molecule composed of one atom of oxygen and two atoms of hydrogen.
2. Water stabilizes body temperature, protects against friction and trauma, makes chemical reactions possible, directly participates in chemical reactions (e.g., dehydration and hydrolysis reactions), and is a mixing medium (e.g., solutions, suspensions, and colloids).
3. A mixture is a combination of two or more substances physically blended together, but not chemically combined.
4. A solution is any liquid, gas, or solid in which the substances are uniformly distributed with no clear boundary between the substances.
5. A solute dissolves in the solvent.
6. A suspension is a mixture containing materials that separate from each other unless they are continually, physically blended together.
7. A colloid is a mixture in which a dispersed (solutelike) substance is distributed throughout a dispersing (solventlike) substance. Particles do not settle out of a colloid.

### Solution Concentrations

1. One way to describe solution concentration is an osmole, which contains $6.022 \times 10^{23}$ of particles (i.e., atoms, ions, or molecules) in 1 kilogram water.
2. A milliosmole is 1/1000 of an osmole.

### Acids and Bases

1. Acids are proton (i.e., hydrogen ion) donors, and bases (e.g., hydroxide ion) are proton acceptors.
2. A strong acid or base almost completely dissociates in water. A weak acid or base partially dissociates.

3. The pH scale refers to the hydrogen ion concentration in a solution.
   - A neutral solution has an equal number of hydrogen ions and hydroxide ions and is assigned a pH of 7.
   - Acid solutions, in which the number of hydrogen ions is greater than the number of hydroxide ions, have pH values less than 7.
   - Basic, or alkaline, solutions have more hydroxide ions than hydrogen ions and a pH greater than 7.
4. A salt is a molecule consisting of a cation other than hydrogen and an anion other than hydroxide. Salts are formed when acids react with bases.
5. A buffer is a solution of a conjugate acid–base pair that resists changes in pH when acids or bases are added to the solution.

### Oxygen

Oxygen is necessary in the reactions that extract energy from food molecules in living organisms.

### Carbon Dioxide

During metabolism when the organic molecules are broken down, carbon dioxide and energy are released.

## Organic Chemistry (p. 43)

Organic molecules contain carbon atoms bound together by covalent bonds.

### Carbohydrates

1. Monosaccharides are the basic building blocks of other carbohydrates. They, especially glucose, are important sources of energy. Examples are ribose, deoxyribose, glucose, fructose, and galactose.
2. Disaccharide molecules are formed by dehydration reactions between two monosaccharides. They are broken apart into monosaccharides by hydrolysis reactions. Examples of disaccharides are sucrose, lactose, and maltose.
3. Polysaccharides are many monosaccharides bound together to form long chains. Examples include cellulose, starch, and glycogen.

### Lipids

1. Triglycerides are composed of glycerol and fatty acids. One, two, or three fatty acids can attach to the glycerol molecule.
   - Fatty acids are straight chains of carbon molecules with a carboxyl group. Fatty acids can be saturated (only single covalent bonds between carbon atoms) or unsaturated (one or more double covalent bonds between carbon atoms).
   - Energy is stored in fats.
2. Phospholipids are lipids in which a fatty acid is replaced by a phosphate-containing molecule. Phospholipids are a major structural component of plasma membranes.
3. Steroids are lipids composed of four interconnected ring molecules. Examples include cholesterol, bile salts, and sex hormones.
4. Other lipids include fat-soluble vitamins, prostaglandins, thromboxanes, and leukotrienes.

### Proteins

1. The building blocks of protein are amino acids, which are joined by peptide bonds.
2. The number, kind, and arrangement of amino acids determine the primary structure of a protein. Hydrogen bonds between amino acids determine secondary structure, and hydrogen bonds between amino acids and water determine tertiary structure. Interactions between different protein subunits determine quaternary structure.
3. Enzymes are protein catalysts that speed up chemical reactions by lowering their activation energy.
4. The active sites of enzymes bind only to specific reactants.
5. Cofactors are ions or organic molecules such as vitamins that are required for some enzymes to function.

### Nucleic Acids: DNA and RNA

1. The basic unit of nucleic acids is the nucleotide, which is a monosaccharide with an attached phosphate and organic base.
2. DNA nucleotides contain the monosaccharide deoxyribose and the organic bases adenine, thymine, guanine, or cytosine. DNA occurs as a double strand of joined nucleotides and is the genetic material of cells.
3. RNA nucleotides are composed of the monosaccharide ribose. The organic bases are the same as for DNA, except that thymine is replaced with uracil.

### Adenosine Triphosphate

ATP stores energy derived from catabolism. The energy is released from ATP and is used in anabolism and other cell processes.

## REVIEW AND COMPREHENSION

1. The smallest particle of an element that still has the chemical characteristics of that element is a (an)
   a. electron.
   b. molecule.
   c. neutron
   d. proton.
   e. atom.
2. The number of electrons in an atom is equal to the
   a. atomic number.
   b. mass number.
   c. number of neutrons.
   d. isotope number.
   e. molecular mass.
3. $^{12}C$ and $^{14}C$ are
   a. atoms of different elements.
   b. isotopes.
   c. atoms with different atomic numbers.
   d. atoms with different numbers of protons.
   e. compounds.
4. A cation is a (an)
   a. uncharged atom.
   b. positively charged atom.
   c. negatively charged atom.
   d. atom that has gained an electron.
   e. both c and d.
5. A polar covalent bond between two atoms occurs when
   a. one atom attracts shared electrons more strongly than another atom.
   b. atoms attract electrons equally.
   c. an electron from one atom is completely transferred to another atom.
   d. the molecule becomes ionized.
   e. a hydrogen atom is shared between two different atoms.
6. Table salt (NaCl) is
   a. an atom.
   b. organic.
   c. a molecule.
   d. a compound.
   e. a cation.

7. The weak attractive force between two water molecules forms a (an)
   a. covalent bond.
   b. hydrogen bond.
   c. ionic bond.
   d. compound.
   e. isotope.
8. Electrolytes are
   a. nonpolar molecules.
   b. covalent compounds.
   c. substances that usually don't dissolve in water.
   d. found in solutions that do not conduct electricity.
   e. cations and anions that dissociate in water.
9. In a decomposition reaction
   a. anabolism occurs.
   b. proteins are formed from amino acids.
   c. large molecules are broken down to form small molecules.
   d. a dehydration reaction may occur.
   e. all of the above.
10. Oxidation–reduction reactions
    a. can be synthesis or decomposition reactions.
    b. have one reactant gaining electrons.
    c. have one reactant losing electrons.
    d. can create ionic or covalent bonds.
    e. all of the above.
11. Potential energy
    a. is energy caused by movement of an object.
    b. is the form of energy that is actually doing work.
    c. includes energy within chemical bonds.
    d. can never be converted to kinetic energy.
    e. all of the above.
12. Which of these descriptions of heat energy is *not* correct?
    a. Heat energy flows between objects that are at different temperatures.
    b. Heat energy can be produced from all other forms of energy.
    c. Heat energy can be released during chemical reactions.
    d. Heat energy must be added to break apart ATP molecules.
    e. Heat energy is always transferred from a hotter object to a cooler object.
13. A *decrease* in the speed of a chemical reaction occurs if
    a. the activation energy requirement is increased.
    b. catalysts are increased.
    c. temperature increases.
    d. the concentration of the reactants increases.
    e. all of the above.
14. Which of these statements concerning enzymes is correct?
    a. Enzymes increase the rate of reactions but are permanently changed as a result.
    b. Enzymes are proteins that function as catalysts.
    c. Enzymes increase the activation energy requirement for a reaction to occur.
    d. Enzymes usually can only double the rate of a chemical reaction.
    e. Enzymes increase the kinetic energy of the reactants.
15. Water
    a. is composed of two oxygen atoms and one hydrogen atom.
    b. has a low specific heat.
    c. is composed of polar molecules into which ionic substances dissociate.
    d. is produced in a hydrolysis reaction.
    e. is a very small organic molecule.
16. When sugar is dissolved in water, the water is called the
    a. solute.
    b. solution.
    c. solvent.
17. Which of these is an example of a suspension?
    a. sweat
    b. water and proteins inside cells
    c. sugar dissolved in water
    d. red blood cells in plasma
18. A solution with a pH of 5 is ________ and contains ________ hydrogen ions than a neutral solution.
    a. a base, more
    b. a base, less
    c. an acid, more
    d. an acid, less
    e. neutral, the same number of
19. A buffer
    a. slows down chemical reactions.
    b. speeds up chemical reactions.
    c. increases the pH of a solution.
    d. maintains a relatively constant pH.
    e. works by forming salts.
20. A conjugate acid–base pair
    a. acts as a buffer.
    b. can combine with hydrogen ions in a solution.
    c. can release hydrogen ions to combine with hydroxide ions.
    d. describes carbonic acid ($H_2CO_3$) and bicarbonate ion ($HCO_3^-$)
    e. all of the above.
21. Carbon dioxide
    a. consists of two oxygen atoms ionically bonded to carbon.
    b. becomes toxic if allowed to accumulate within cells.
    c. is mostly eliminated by the kidneys.
    d. is combined with fats to produce glucose during metabolism within cells.
    e. is taken into cells during metabolism.
22. Which of these is an example of a carbohydrate?
    a. glycogen
    b. prostaglandin
    c. steroid
    d. DNA
    e. triglyceride
23. The polysaccharide used for energy storage in the human body is
    a. cellulose.
    b. glycogen.
    c. lactose.
    d. sucrose.
    e. starch.
24. The basic units or building blocks of triglycerides are
    a. simple sugars (monosaccharides).
    b. double sugars (disaccharides).
    c. amino acids.
    d. glycerol and fatty acids.
    e. nucleotides.
25. A ________ fatty acid has one double covalent bond between carbon atoms.
    a. cholesterol
    b. monounsaturated
    c. phospholipid
    d. polyunsaturated
    e. saturated
26. A peptide bond joins together
    a. amino acids.
    b. fatty acids and glycerol.
    c. monosaccharides.
    d. disaccharides.
    e. nucleotides.
27. The ________ structure of a protein results from the folding of the helices or pleated sheets.
    a. primary
    b. secondary
    c. tertiary
    d. quaternary

28. According to the lock-and-key model of enzyme action,
   a. reactants must first be heated.
   b. enzyme shape is not important.
   c. each enzyme can catalyze many types of reactions.
   d. reactants must bind to an active site on the enzyme.
   e. enzymes control only a small number of reactions in the cell.
29. DNA molecules
   a. are the genetic material of cells.
   b. contain a single strand of nucleotides.
   c. contain the nucleotide uracil.
   d. have three different types that have roles in protein synthesis.
   e. contain up to 100 organic bases.
30. ATP
   a. is formed by the addition of a phosphate group to ADP.
   b. is formed with energy released during catabolism reactions.
   c. provides the energy for anabolism reactions.
   d. contains three phosphate groups.
   e. all of the above.

*Answers in Appendix F*

## C R I T I C A L T H I N K I N G

1. Iron has an atomic number of 26 and a mass number of 56. How many protons, neutrons, and electrons are in an atom of iron? If an atom of iron lost three electrons, what would the charge of the resulting ion be? Write the correct symbol for this ion.
2. Which of the following pairs of terms applies to the reaction that results in the formation of fatty acids and glycerol from a triglyceride molecule?
   a. Decomposition or synthesis reaction
   b. Anabolism or catabolism
   c. Dehydration or hydrolysis reaction
3. A mixture of chemicals is warmed slightly. As a consequence, although no more heat is added, the solution becomes very hot. Explain what occurred to make the solution so hot.
4. Two solutions, when mixed together at room temperature, produce a chemical reaction. When the solutions are boiled and allowed to cool to room temperature before mixing, however, no chemical reaction takes place. Explain.
5. In terms of the potential energy in the food, explain why eating food is necessary for increasing muscle mass.
6. Solution A has a pH of 2, and solution B has a pH of 8. If equal amounts of solutions A and B are mixed, is the resulting solution acidic or basic?
7. Given a buffered solution that is based on the following equilibrium:
$$CO_2 + H_2O \rightleftarrows H_2CO_3 \rightleftarrows H^+ + HCO_3^-$$
what happens to the pH of the solution if $NaHCO_3$ is added?
8. An enzyme E catalyzes the following reaction:
$$A + B \xrightarrow{E} C$$
The product C, however, binds to the active site of the enzyme in a reversible fashion and keeps the enzyme from functioning. What happens if A and B are continually added to a solution that contains a fixed amount of the enzyme?
9. Given the materials commonly found in a kitchen, explain how one could distinguish between a protein and a lipid.
10. A student is given two unlabeled substances: one a typical phospholipid and one a typical protein. She is asked to determine which substance is the protein and which is the phospholipid. The available techniques allow her to determine the elements in each sample. How can she identify each substance?

*Answers in Appendix G*

## A N S W E R S T O P R E D I C T Q U E S T I O N S

1. The mass (amount of matter) of the astronaut on the surface of the earth and in outer space does not change. In outer space, where the force of gravity from the earth is very small, the astronaut is "weightless" compared to his or her weight on the earth's surface.
2. Potassium has 19 protons (the atomic number), 20 neutrons (the mass number minus the atomic number), and 19 electrons (because the number of electrons equals the number of protons).
3. The molecular formula for glucose is $C_6H_{12}O_6$. The atomic mass of carbon is 12.01, hydrogen is 1.008, and oxygen is 16.00. The molecular mass of glucose is therefore $(6 \times 12.01) + (12 \times 1.008) + (6 \times 16.00)$, or 180.2.
4. A decrease in blood $CO_2$ decreases the amount of $H_2CO_3$ and therefore the blood $H^+$ level. Because $CO_2$ and $H_2O$ are in equilibrium with $H^+$ and $HCO_3^-$, with $H_2CO_3$ as an intermediate, a decrease in $CO_2$ causes some $H^+$ and $HCO_3^-$ to join together to form $H_2CO_3$, which then forms $CO_2$ and $H_2O$. Consequently, the $H^+$ concentration decreases.
5. When two hydrogen atoms combine with an oxygen atom to form water, a polar covalent bond forms between each hydrogen atom and the oxygen atom. Unequal sharing of electrons occurs, and the electrons are associated with the oxygen atom more than with the hydrogen atoms. In this sense, the hydrogen atoms lose their electrons, and the oxygen atom gains electrons. The hydrogen atoms are therefore oxidized, and the oxygen atom is reduced.
6. During exercise, muscle contractions increase, which requires energy. This energy is obtained from the energy in the chemical bonds of ATP. As ATP is broken down, energy is released. Some of the energy is used to drive muscle contractions, and some becomes heat. Because the rate of these reactions increases during exercise, more heat is produced than when at rest, and body temperature increases.
7. Monohydrogen phosphate ion ($HPO_4^{2-}$) is the conjugate base formed when the conjugate acid, dihydrogen phosphate ion ($H_2PO_4^-$) loses a hydrogen ion. If hydrogen ions are added to the solution, they combine with the conjugate base, monohydrogen phosphate ions, to form dihydrogen phosphate ions, which helps to prevent an increase in hydrogen ion concentration. If hydroxide ions are added to the solution, they combine with hydrogen ions to form water. Then the conjugate acid, dihydrogen phosphate ions, dissociate to replace the hydrogen ions, which helps to prevent a decrease in hydrogen ion concentration.

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