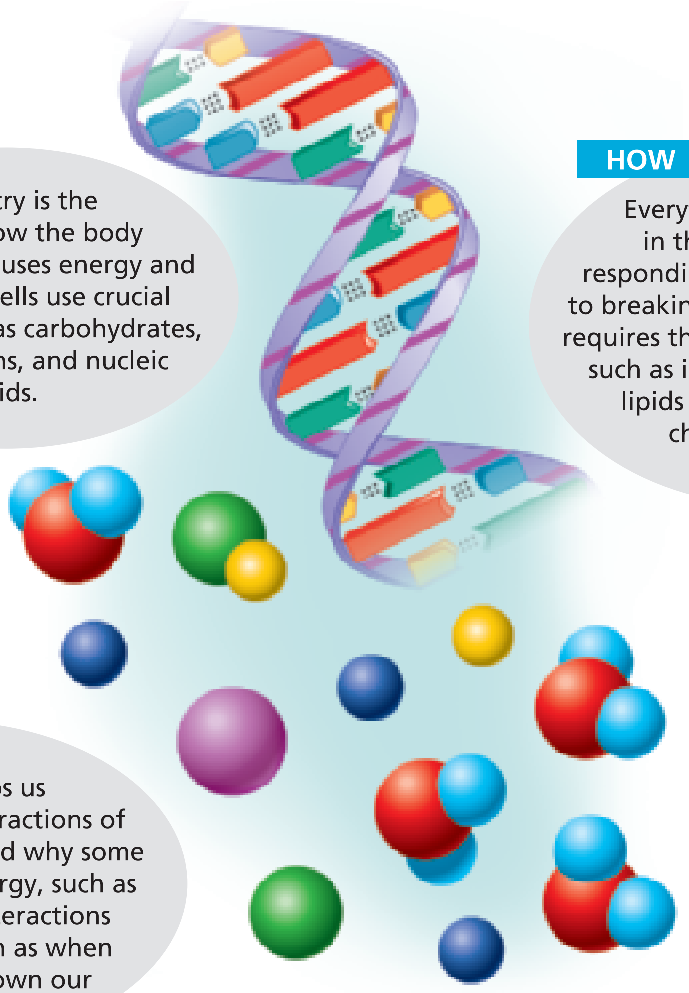
2 Basic Chemistry



WHAT

Chemistry is the basis for how the body transforms and uses energy and for how our cells use crucial molecules such as carbohydrates, lipids, proteins, and nucleic acids.

HOW

Everything that happens in the body, from cells responding to their surroundings to breaking down the food we eat, requires the movement of chemicals such as ions, carbohydrates, and lipids as they participate in chemical reactions.

WHY

Chemistry helps us understand the interactions of different molecules and why some interactions store energy, such as in fat, and other interactions release energy, such as when enzymes break down our food.

INSTRUCTORS

New **Building Vocabulary Coaching** Activities for this chapter are assignable in MasteringA&P®

Many short courses in anatomy and physiology lack the time to consider chemistry as a topic. So why include it here? The answer is simple: Your entire body is made up of chemicals—thousands of them—continuously interacting with one another at an incredible pace.

Chemical reactions underlie all body processes—movement, digestion, the pumping of your heart, and even your thoughts. In this chapter we present the basics of chemistry and biochemistry (the chemistry of living material), providing the background you will need to understand body functions.

Concepts of Matter and Energy

→ Learning Objectives

- ☐ **Differentiate matter from energy.**
- ☐ **List four major energy forms, and provide one example of how each is used in the body.**

Matter

Matter is the "stuff" of the universe. With some exceptions, it can be seen, smelled, and felt. More precisely, matter is anything that occupies space and has mass. *Weight* is a measure of gravity pulling on mass. Chemistry studies the nature of matter—how its building blocks are put together and how they interact.

Matter exists in solid, liquid, and gaseous states, all of which are found in the human body. *Solids*, such as bones and teeth, have a definite shape and volume. *Liquids* have a definite volume, but they conform to the shape of their container. Examples of body liquids are blood plasma and the interstitial fluid that bathes all body cells. *Gases* have neither a definite shape nor a definite volume. The air we breathe is a mixture of gases.

Matter may be changed both physically and chemically. *Physical changes* do not alter the basic nature of a substance. Examples include changes in state, such as ice melting to water and food being cut into smaller pieces. *Chemical changes* do alter the composition of the substance—often substantially. Fermenting grapes to make wine and the digestion of food in the body are examples of chemical changes.

Energy

In contrast to matter, **energy** has no mass and does not take up space. It can be measured only by its effects on matter. We commonly define energy as the ability to do work or to put matter into motion. When energy is actually doing work (moving objects), we refer to it as **kinetic** (kĭ-neh′tik) **energy**. Kinetic energy is displayed in the constant movement of the tiniest particles of matter (atoms) as well as in larger objects, such as a bouncing ball. When energy is inactive or stored (as in the batteries of an unused toy), we call it **potential energy**. All forms of energy exhibit both kinetic and potential work capacities.

All living things are built of matter, and to grow and function they require a continuous supply of energy. Thus, matter is the substance, and energy is the mover of the substance. Because this is so, let's take a brief detour to introduce the forms of energy the body uses as it does its work.

Forms of Energy

- **Chemical energy** is stored in the bonds of chemical substances. When the bonds are broken, the (potential) stored energy is unleashed and becomes kinetic energy (energy in action). For example, when gasoline molecules are broken apart in your automobile engine, the energy released powers your car. Similarly, the chemical energy harvested from the foods we eat fuels all body activities.
- **Electrical energy** results from the movement of charged particles. In your house, electrical energy is the flow of subatomic particles called electrons along the wiring. In your body, an electrical current is generated when charged particles (called *ions*) move across cell membranes. The nervous system uses electrical currents called *nerve impulses* to transmit messages from one part of the body to another.
- **Mechanical energy** is energy *directly* involved in moving matter. When you pedal a bicycle, your legs provide the mechanical energy that turns the wheels. We can take this example one step further back: As the muscles in your legs shorten (contract), they pull on your bones, causing your limbs to move (so that you can pedal the bike).
- **Radiant energy** travels in waves; that is, it is the energy of the electromagnetic spectrum, which includes X rays, infrared radiation (heat energy), visible light, radio, and ultraviolet (UV) waves. Light energy, which stimulates the retinas of your eyes, is important in vision. UV waves cause sunburn, but they also stimulate our bodies to make vitamin D.

Energy Form Conversions

With a few exceptions, energy is easily converted from one form to another. For example, chemical energy (from gasoline) that powers the motor of a speedboat is converted into the mechanical energy of the whirling propeller that allows the boat to skim across the water. In the body, chemical energy from food is trapped in the bonds of a high-energy chemical called **ATP (adenosine triphosphate)**, and ATP's energy may ultimately be transformed into the electrical energy of a nerve impulse or the mechanical energy of contracting muscles.

Energy conversions are not very efficient, and some of the initial energy supply is always "lost" to the environment as heat (thermal energy). It is not really lost, because energy cannot be created or destroyed, but the part given off as heat is *unusable*. You can easily demonstrate this principle by putting your finger close to a lightbulb that has been lit for an hour or so. Notice that some of the electrical energy reaching the bulb is producing heat instead of light. Likewise, all energy conversions in the body liberate heat. This heat makes us warm-blooded animals and contributes to our relatively high body temperature, which has an important influence on body function. For example, when matter is heated, its particles begin to move more quickly; that is, their kinetic energy (energy of motion) increases. This is important to the chemical reactions that occur in the body because, up to a point, the higher the temperature, the faster those reactions occur.

Did You Get It?

1. Matter and energy—how are they interrelated?
2. What form of energy is used to transmit messages from one part of the body to another? What form of energy fuels cellular processes?
3. What type of energy is available when we are still? When we are exercising?
4. What does it mean when we say that some energy is "lost" every time energy changes from one form to another in the body?

For answers, see Appendix A.

Composition of Matter

Elements and Atoms

➔ Learning Objectives

- ☐ **Define *element*, and list the four elements that form the bulk of body matter.**
- ☐ **Explain how elements and atoms are related.**

All matter is composed of a limited number of substances called **elements**, unique substances that cannot be broken down into simpler substances by ordinary chemical methods. Examples of elements include many commonly known substances, such as oxygen, silver, gold, copper, and iron.

So far, 118 elements have been identified with certainty. Ninety-two of these occur in nature; the rest are made artificially in accelerator devices. A complete listing of the elements appears in the **periodic table**, an odd-shaped checkerboard (see Appendix C) that appears in chemistry classrooms the world over. Four of these elements—carbon, oxygen, hydrogen, and nitrogen—make up about 96 percent of the weight of the human body. Besides these four elements, several others are present in small or trace amounts. The most abundant elements found in the body and their major roles are listed in Table 2.1.

Each element is composed of very similar particles, or building blocks, called **atoms**. Because all elements are unique, the atoms of each element differ from those of all other elements. We designate each element by a one- or two-letter chemical shorthand called an **atomic symbol**. In most cases, the atomic symbol is simply the first or first two letters of the element's name. For example, C stands for carbon, O for oxygen, and Ca for calcium. In a few cases, the atomic symbol is taken from the Latin name for the element. For instance, Na (from the Latin word *natrium*) indicates sodium, and K (from *kalium*) potassium.

Atomic Structure

➔ Learning Objective

- ☐ **List the subatomic particles, and describe their relative masses, charges, and positions in the atom.**

The Basic Atomic Subparticles

The word *atom* comes from the Greek word meaning "incapable of being divided," and historically this idea of an atom was accepted as a scientific truth. According to this notion, you could theoretically divide a pure element, such as a block of gold, into smaller and smaller particles until you got down to the individual atoms, at which point you could subdivide no further. We now know that atoms, although indescribably small, are clusters of even smaller components called **subatomic particles**, which include *protons*, *neutrons*, and *electrons*. Even so, the old idea of atomic indivisibility is still very useful because an atom loses the unique properties of its element when it is split into its component particles.

An atom's subatomic particles differ in their mass, electrical charge, and position within the atom (Table 2.2, p. 28). **Protons (p^+)** have a positive charge, whereas **neutrons (n^0)** are uncharged, or neutral. Protons and neutrons are heavy particles and have approximately the same mass (1 atomic mass unit, or amu). The tiny **electrons (e^-)** bear a negative charge equal in strength to the positive charge of the protons, but their mass is so small that it is usually designated as 0 amu.

The electrical charge of a particle is a measure of its ability to attract or repel other charged particles. Particles with the same type of charge (+ to + or – to –) repel each other, but particles with opposite charges (+ to –) attract each other.

Table 2.1 Common Elements Making Up the Human Body

Element	Atomic symbol	Percentage of body mass	Role
Major (96.1%)			
Oxygen	O	65.0	A major component of both organic and inorganic molecules; as a gas, essential to the oxidation of glucose and other food fuels, during which cellular energy (ATP) is produced.
Carbon	C	18.5	The primary element in all organic molecules, including carbohydrates, lipids, proteins, and nucleic acids.
Hydrogen	H	9.5	A component of most organic molecules; as an ion (a charged atom), it influences the pH of body fluids.
Nitrogen	N	3.2	A component of proteins and nucleic acids (genetic material).
Lesser (3.9%)			
Calcium	Ca	1.5	Found as a salt in bones and teeth; in ionic form, required for muscle contraction, neural transmission, and blood clotting.
Phosphorus	P	1.0	Present as a salt, in combination with calcium, in bones and teeth; also present in nucleic acids and many proteins; forms part of the high-energy compound ATP.
Potassium	K	0.4	In its ionic form, the major intracellular cation; necessary for the conduction of nerve impulses and for muscle contraction.
Sulfur	S	0.3	A component of proteins (particularly contractile proteins of muscle).
Sodium	Na	0.2	As an ion, the major extracellular cation; important for water balance, conduction of nerve impulses, and muscle contraction.
Chlorine	Cl	0.2	In ionic (chloride) form, the most abundant extracellular anion.
Magnesium	Mg	0.1	Present in bone; also an important cofactor for enzyme activity in a number of metabolic reactions.
Iodine	I	0.1	Needed to make functional thyroid hormones.
Iron	Fe	0.1	A component of the functional hemoglobin molecule (which transports oxygen within red blood cells) and some enzymes.
Trace (less than 0.01%)*			
Chromium (Cr), Cobalt (Co), Copper (Cu), Fluorine (F), Manganese (Mn), Molybdenum (Mo), Selenium (Se), Silicon (Si), Tin (Sn), Vanadium (V), Zinc (Zn)			

*Referred to as the *trace elements* because they are required in very small amounts; many are found as part of enzymes or are required for enzyme activation.

Neutral particles are neither attracted to nor repelled by charged particles.

Because all atoms are electrically neutral, the number of protons an atom has must be balanced by its number of electrons (the + and – charges will then cancel the effect of each other). Thus, helium has two protons and two electrons, and iron has 26 protons and 26 electrons. For any atom, the number of protons and electrons is always equal. Atoms that have gained or lost electrons are called *ions*.

Table 2.2 Subatomic Particles

Particle	Position in atom	Mass (amu)	Charge
Proton (p^+)	Nucleus	1	+
Neutron (n^0)	Nucleus	1	0
Electron (e^-)	Orbits around the nucleus	1/2000*	–

*The mass of an electron is so small, that we will ignore it and assume a mass of 0 amu.

Planetary and Orbital Models of an Atom

The **planetary model** of an atom portrays the atom as a miniature solar system **(Figure 2.1a)** in which the protons and neutrons are clustered at the center of the atom in the **atomic nucleus**. Because the nucleus contains all the heavy particles, it is fantastically dense and positively charged. The tiny electrons orbit around the nucleus in fixed, generally circular orbits, like planets around the sun. But we can never determine the exact location of electrons at a particular time because they jump around following unknown paths. So, instead of speaking of specific orbits, chemists talk about *orbitals*—regions around the nucleus in which electrons are *likely* to be found. The **orbital model** depicts the general location of electrons outside the nucleus as a haze of negative charge referred to as the *electron cloud* (Figure 2.1b). Regions where electrons are most likely to be found are shown by denser shading rather than by orbit lines.

Notice that in both models, the electrons have the run of nearly the entire volume of the atom. Electrons also determine an atom's chemical behavior (that is, its ability to bond with other atoms). Though now considered outdated, the planetary model is simple and easy to understand and use. Most of the descriptions of atomic structure in this text use that model.

Hydrogen is the simplest atom, with just one proton and one electron. You can visualize the hydrogen atom by imagining it as a sphere with its diameter equal to the length of a football field. The nucleus could then be represented by a lead ball the size of a gumdrop in the exact center of the sphere and the lone electron pictured as a fly buzzing about unpredictably within the sphere. This mental picture should remind you that most of the volume of an atom is empty space, and most of the mass is in the central nucleus.

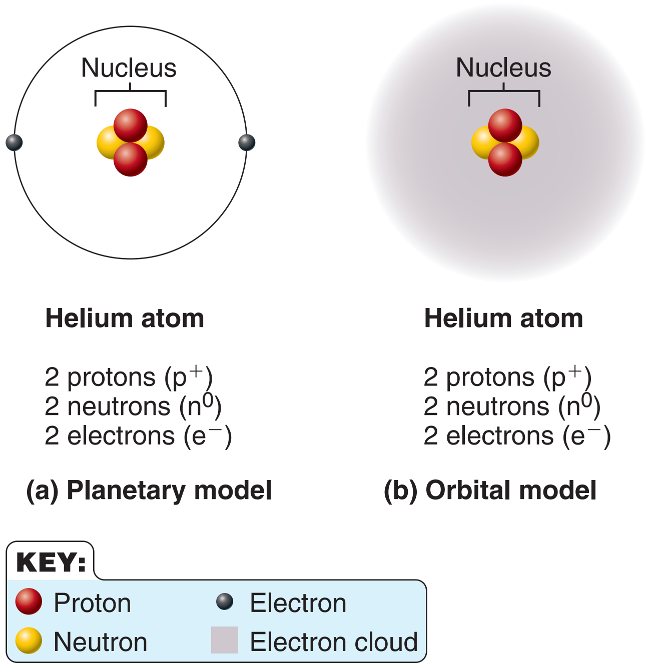


Figure 2.1 The structure of an atom. The dense central nucleus contains the protons and neutrons. **(a)** In the planetary model of an atom, the electrons move around the nucleus in fixed orbits. **(b)** In the orbital model, electrons are shown as a cloud of negative charge.

Did You Get It?

5. Which four elements make up the bulk of living matter?
6. How is an atom related to an element?

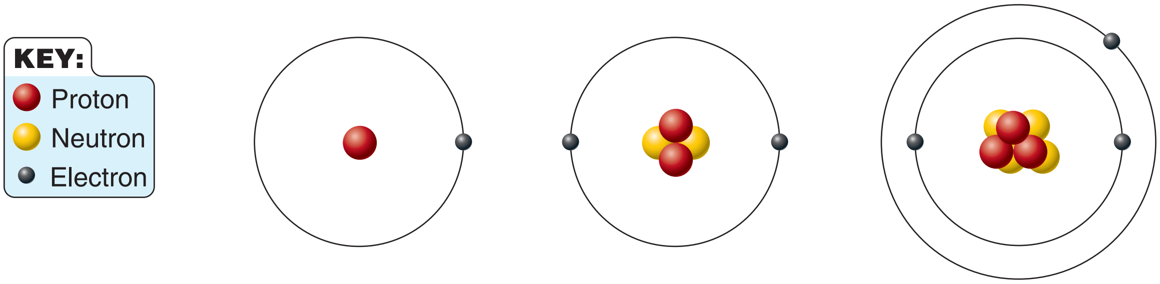
For answers, see Appendix A.

Identifying Elements

→ Learning Objective

☐ **Define *radioisotope*, and describe briefly how radioisotopes are used in diagnosing and treating disease.**

All protons are alike, regardless of the atom being considered. The same is true of all neutrons and all electrons. So what determines the unique properties



(a) Hydrogen (H) ($1p^+$; $0n^0$; $1e^-$)

(b) Helium (He) ($2p^+$; $2n^0$; $2e^-$)

(c) Lithium (Li) ($3p^+$; $4n^0$; $3e^-$)

Figure 2.2 Atomic structure of the three smallest atoms.

of each element? The answer is that atoms of different elements are composed of *different numbers* of protons, neutrons, and electrons.

The simplest and smallest atom, hydrogen, has one proton, one electron, and no neutrons **(Figure 2.2)**. Next is the helium atom, with two protons, two neutrons, and two orbiting electrons. Lithium follows with three protons, four neutrons, and three electrons. If we continue this step-by-step listing of subatomic particles, we could describe all known atoms by adding one proton and one electron at each step. The number of neutrons is not as easy to pin down, but light atoms tend to have equal numbers of protons and neutrons, whereas in larger atoms neutrons outnumber protons. However, all we really need to know to identify a particular element is its *atomic number*, *atomic mass number*, and *atomic weight*. Taken together, these indicators provide a fairly complete picture of each element.

Atomic Number

Each element is given a number, called its **atomic number**, that is equal to the number of protons its atoms contain. Atoms of each element contain a different number of protons from the atoms of any other element; hence, its atomic number is unique. Because the number of protons is always equal to the number of electrons, the atomic number *indirectly* also tells us the number of electrons that atom contains.

Atomic Mass Number

The **atomic mass number** (or just **mass number**) of any atom is the sum of the masses of all the protons and neutrons contained in its nucleus. (Remember, the mass of an electron is so small that we ignore it.) Because hydrogen has one proton and no neutrons in its nucleus, its atomic number and mass number are the same: 1. Helium, with two protons and two neutrons, has a mass number of 4. The mass number is written as a superscript to the left of the atomic symbol (see the examples in **Figure 2.3** on p. 30).

Atomic Weight and Isotopes

At first glance, it seems that the **atomic weight** of an atom should be equal to its atomic mass. This would be so if there were only one type of atom representing each element. However, the atoms of almost all elements exhibit two or more structural variations; these varieties are called **isotopes** (iʹ sĭ-tōps). Isotopes have the same number of protons and electrons but vary in the number of *neutrons* they contain. Thus, the isotopes of an element have the same atomic number but different atomic masses. Because all of an element's isotopes have the same number of electrons (and electrons determine bonding properties), their chemical properties are *exactly* the same. As a general rule, the atomic weight of any element is approximately equal to the mass number of its most abundant isotope. For example, as we said before, hydrogen has an atomic number of 1, but it also has isotopes with atomic masses of 1, 2 (called *deuterium*), and 3 (called *tritium*) (see Figure 2.3). Its atomic weight is 1.008, which reveals that its lightest isotope is present in much greater amounts in our world than its ^2H or ^3H forms. (The atomic numbers, mass numbers, and atomic weights for elements commonly found in the body are provided in **Table 2.3**).

Q: ***Which of these isotopes is the heaviest?***

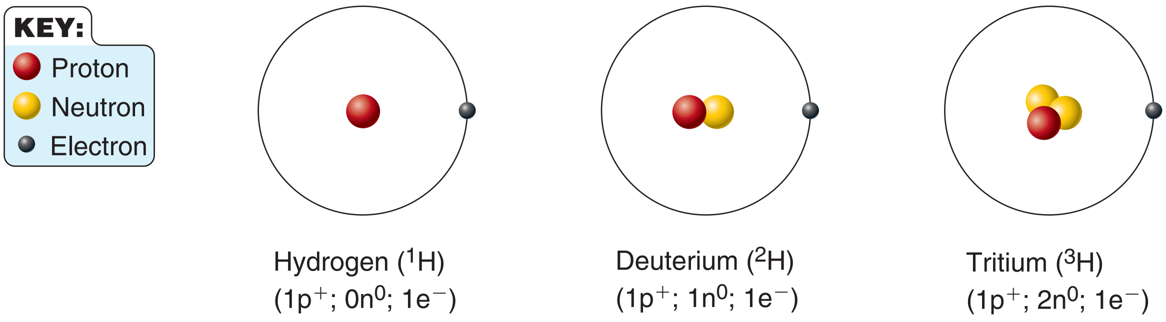


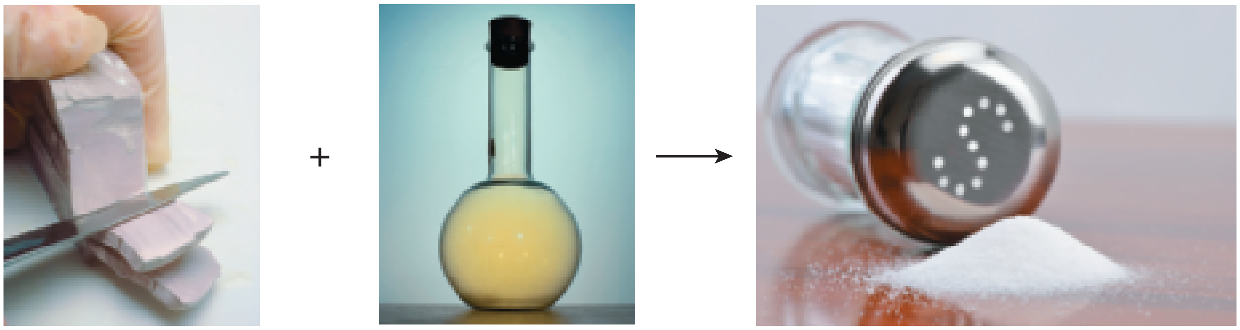
Figure 2.3 Isotopes of hydrogen. Isotopes differ in their numbers of neutrons.

The heavier isotopes of certain atoms are unstable and tend to release energy in order to become more stable; such isotopes are called **radioisotopes**. Why this process occurs is very complex, but apparently the “glue” that holds the atomic nuclei together is weaker in the heavier isotopes. This process of spontaneous atomic decay, which is called **radioactivity**, can be compared to a tiny sustained explosion. All types of radioactive decay involve the ejection of particles (*alpha* or *beta particles*) or electromagnetic energy (*gamma rays*) from the atom’s nucleus and are damaging to living cells. Emission of alpha particles has the least penetrating power; gamma

Table 2.3 Atomic Structures of the Most Abundant Elements in the Body

Element	Symbol	Atomic number (# of p)	Mass number (# of p + n)	Atomic weight	Electrons in valence shell
Calcium	Ca	20	40	40.078	2
Carbon	C	6	12	12.011	4
Chlorine	Cl	17	35	35.453	7
Hydrogen	H	1	1	1.008	1
Iodine	I	53	127	126.905	7
Iron	Fe	26	56	55.847	2
Magnesium	Mg	12	24	24.305	2
Nitrogen	N	7	14	14.007	5
Oxygen	O	8	16	15.999	6
Phosphorus	P	15	31	30.974	5
Sodium	Na	11	23	22.989	1
Sulfur	S	16	32	32.064	6

A: Tritium.



Sodium (silvery metal) + Chlorine (poisonous gas) → Sodium chloride (table salt)

Figure 2.4 Properties of a compound differ from those of its atoms.

radiation has the most. Contrary to what some believe, ionizing radiation does not damage the atoms in its path directly. Instead, it sends electrons flying, like a bowling ball through pins, all along its path. It is these electrons that do the damage.

Radioisotopes are used in minute amounts to tag biological molecules so that they can be followed, or traced, through the body and are valuable tools for medical diagnosis and treatment. For example, PET scans, which use radioisotopes, are discussed in "A Closer Look" (pp. 10–11). A radioisotope of iodine can be used to scan the thyroid gland of a patient suspected of having a thyroid tumor. Additionally, radium, cobalt, and certain other radioisotopes are used to destroy localized cancers.

Did You Get It?

7. An atom has five neutrons, four protons, and four electrons. What is its atomic number? What is its atomic mass number?
8. What name is given to an unstable atom that has either more or fewer neutrons than its typical number?

For answers, see Appendix A.

Molecules and Compounds

→ Learning Objective

☐ **Define *molecule*, and explain how molecules are related to compounds.**

When two or more atoms combine chemically, **molecules** are formed. If two or more atoms of the same element bond, or become chemically linked together, a molecule of that element is produced. For example, when two hydrogen atoms bond, a molecule of hydrogen gas is formed:

$$\text{H (atom)} + \text{H (atom)} \rightarrow \text{H}_2 \text{ (molecule)}^*$$

In this example of a chemical reaction, the *reactants* (the atoms taking part in the reaction) are indicated by their atomic symbols, and the *product* (the molecule formed) is indicated by a *molecular formula* that shows its atomic makeup. The chemical reaction is shown as a *chemical equation*.

When two or more *different* atoms bind together to form a molecule, the molecule is more specifically referred to as a molecule of a **compound**. For example, four hydrogen atoms and one carbon atom can interact chemically to form methane:

$$4\text{H} + \text{C} = \text{CH}_4 \text{ (methane)}$$

Thus, a molecule of methane is a compound, but a molecule of hydrogen gas is not—it is instead called molecular hydrogen.

Compounds always have properties quite different from those of the atoms making them up, and it would be next to impossible to determine the atoms making up a compound without analyzing it chemically. Sodium chloride is an excellent example of the difference in properties between a compound and its constituent atoms (Figure 2.4). Sodium is a silvery white metal, and chlorine in its

*Notice that when the number of atoms is written as a subscript, it indicates that the atoms are joined by a chemical bond. Thus, 2H represents two separate atoms, but H_2 indicates that the two hydrogen atoms are bound together to form a molecule. The atomic symbol by itself represents one atom.

molecular state is a poisonous green gas used to make bleach. However, sodium chloride is table salt, a white crystalline solid that we sprinkle on our food. Notice that just as an atom is the smallest particle of an element that still retains that element's properties, a molecule is the smallest particle of a compound that still retains the properties of that compound. If you break the bonds between the atoms of the compound, properties of the atoms, rather than those of the compound, will be exhibited.

Did You Get It?

9. What is the meaning of the term molecule?
10. How does a molecular substance differ from a molecule of a compound?

For answers, see Appendix A.

Chemical Bonds and Chemical Reactions

➔ Learning Objectives

- ☐ **Recognize that chemical reactions involve the interaction of electrons to make and break chemical bonds.**
- ☐ **Differentiate between ionic, polar covalent, and nonpolar covalent bonds, and describe the importance of hydrogen bonds.**
- ☐ **Contrast synthesis, decomposition, and exchange reactions.**

Chemical reactions occur whenever atoms combine with or dissociate from other atoms. When atoms unite chemically, chemical bonds are formed.

Bond Formation

A **chemical bond** is not an actual physical structure, like a pair of handcuffs linking two people together. Instead, it is an energy relationship that involves interactions between the electrons of the reacting atoms. Let's consider the role electrons play in forming bonds.

Role of Electrons

The orbits, or generally fixed regions of space that electrons occupy around the nucleus (see Figure 2.2), are called **electron shells**, or **energy levels**. The maximum number of electron shells in any atom known so far is seven, and these are numbered 1 to 7 from the nucleus outward. The electrons closest to the nucleus are those most strongly attracted to its positive charge, and those farther away are less securely held. As a result, the more distant electrons are likely to interact with other atoms. Put more simply, the electron shell furthest from the nucleus is also the first part of the atom any other atom will come into contact with prior to reacting or bonding.

There is an upper limit to the number of electrons that each electron shell can hold. Shell 1, closest to the nucleus, is small and can accommodate only 2 electrons. Shell 2 holds a maximum of 8. Shell 3 can accommodate up to 18 electrons. Subsequent shells hold larger and larger numbers of electrons. In most (but not all) cases, the shells tend to be filled consecutively.

The only electrons that are important to bonding behavior are those in the atom's outermost shell. This shell is called the **valence shell**, and its *valence electrons* determine the chemical behavior of the atom. As a general rule, the electrons of inner shells do not take part in bonding.

The key to chemical reactivity is referred to as the *rule of eights*: In the absence of a full valence shell, atoms interact in such a way that they will have eight electrons in their valence shell. The first electron shell represents an exception to this rule, because it is "full" when it has two electrons. When the valence shell of an atom contains eight electrons, the atom is completely stable and is chemically inactive (inert). When the valence shell contains fewer than eight electrons, an atom will tend to gain, lose, or share electrons with other atoms to reach a stable state. When any of these events occurs, chemical bonds are formed. (Examples of chemically inert and reactive elements are shown in Figure 2.5). As you might guess, atoms must approach each other very closely for their electrons to interact in a chemical bond—in fact, their outermost electron shells must overlap.

Types of Chemical Bonds

Ionic Bonds **Ionic** (i-on′ik) **bonds** form when electrons are completely transferred from one atom to another. Atoms are electrically neutral, but when they gain or lose electrons during bonding, their positive and negative charges are no longer balanced, and charged particles, called **ions**, result. When an atom gains an electron, it acquires a net negative charge because it now has more electrons than protons. Negatively charged ions are more specifically called *anions*, and they are

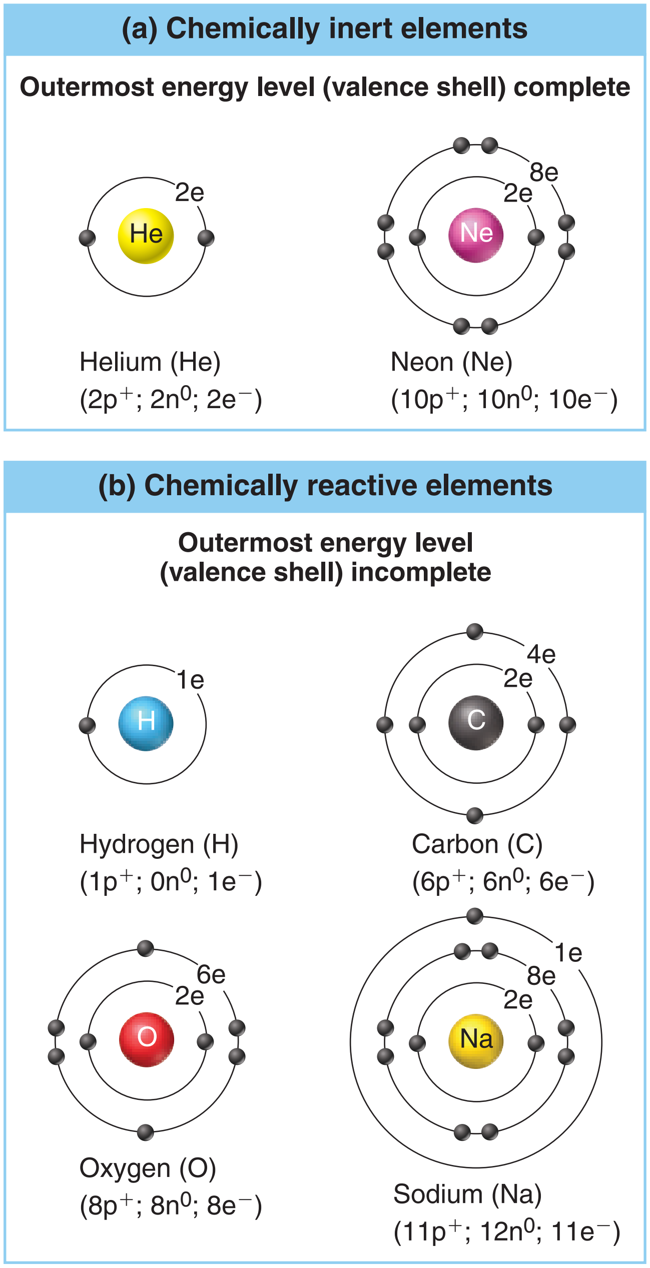


Figure 2.5 Chemically inert and reactive elements. To simplify the diagrams, each atomic nucleus is shown as a circle with the atom's symbol in it; protons and neutrons are not shown.

indicated by a minus sign after the atomic symbol, such as Cl^- for the chloride ion. When atoms lose an electron, they become positively charged ions, *cations*, because they now possess more protons than electrons. Cations are represented by their atomic symbol with a plus sign, such as H^+ for the hydrogen ion (it may help you to remember that a cation is positively charged by thinking of its "t" as a plus [+] sign). Both anions and cations result when an ionic bond is formed. Because opposite charges attract, the newly created ions tend to stay close together.

The formation of sodium chloride (NaCl), common table salt, provides a good example of ionic bonding. Sodium's valence shell contains only one electron and so is incomplete (**Figure 2.6**, p. 34). However, if this single electron is "lost" to another atom, shell 2, which contains eight electrons, becomes the valence shell; thus sodium becomes a cation (Na^+) and achieves stability. Chlorine needs only one electron to fill its valence shell, and it is much easier to gain one electron (forming Cl^-) than it is to "give away" seven. Thus, the ideal situation is for sodium to donate its valence electron to chlorine, which is exactly what happens. Sodium chloride and most other compounds formed by ionic bonding fall into the general category of inorganic chemicals called *salts*.

Covalent Bonds Electrons do not have to be completely lost or gained for atoms to become stable. Instead, they can be shared in such a way that each atom is able to fill its valence shell at least part of the time.

Molecules in which atoms share valence electrons are called *covalent molecules*, and their bonds are **covalent bonds** (*co* = with; *valent* = having power). For example, hydrogen, with its single electron, can become stable if it fills its valence shell (energy level 1) by sharing a pair of electrons—its own and one from another atom. A hydrogen atom can share an electron pair with another hydrogen atom to form a molecule of hydrogen gas **(Figure 2.7a)**. The shared electron pair orbits the whole molecule and satisfies the stability needs of both hydrogen atoms. Likewise, two oxygen atoms, each with six valence electrons, can share two pairs of electrons (form double bonds) with each other (Figure 2.7b) to form a molecule of oxygen gas (O_2).

A hydrogen atom may also share its electron with an atom of a different element. Carbon has four valence electrons but needs eight to achieve stability. When methane gas (CH_4) is formed, carbon shares four electrons with four hydrogen atoms (each bond includes one electron from each atom to form a pair of electrons; Figure 2.7c). Because the shared electrons orbit and "belong to"

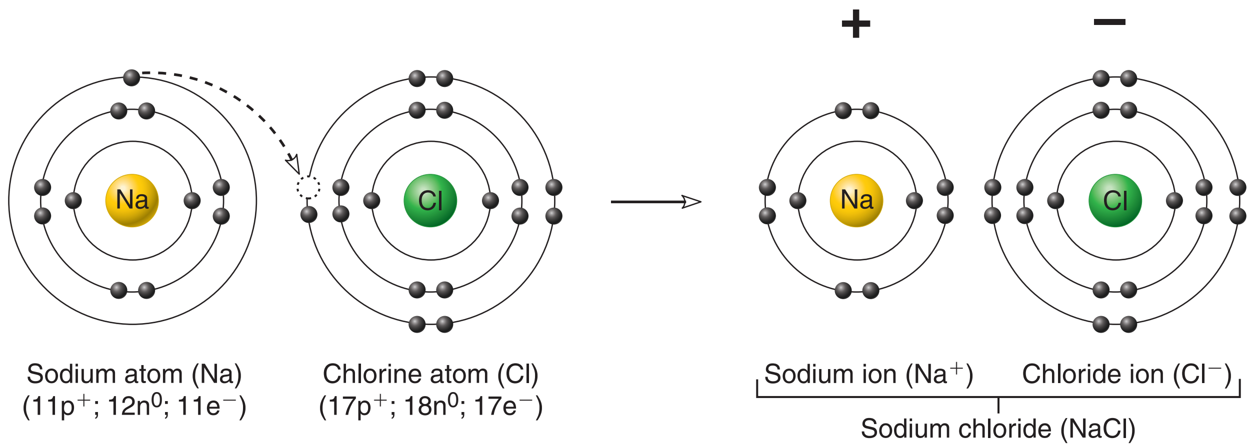


Figure 2.6 Formation of an ionic bond. Both sodium and chlorine atoms are chemically reactive because their valence shells are incompletely filled. Sodium gains stability by losing one electron, whereas chlorine becomes stable by gaining one electron. After electron transfer, sodium becomes a sodium ion (Na^+), and chlorine becomes a chloride ion (Cl^-). The oppositely charged ions attract each other.

the whole molecule, each atom has a full valence shell enough of the time to satisfy its stability needs.

In the covalent molecules described thus far, electrons have been shared *equally* between the atoms of the molecule. Such molecules are called *nonpolar covalent molecules*. However, electrons are not in all cases shared equally. When covalent bonds are made, the molecule formed always has a definite three-dimensional shape. A molecule's shape plays a major role in determining just what other molecules (or atoms) it can interact with; the shape may also result in unequal electron-pair sharing. The following two examples illustrate this principle.

Carbon dioxide (CO_2) is formed when a carbon atom shares its four valence electrons with two oxygen atoms. Oxygen is a very "electron-hungry" atom, so it attracts the shared electrons much more strongly than does carbon. However, because the carbon dioxide molecule is linear (O═C═O), the electron-pulling power of one oxygen atom is offset by that of the other, like a tug-of-war at a standoff between equally strong teams (**Figure 2.8a**, p. 36). As a result, the electron pairs are shared equally and orbit the entire molecule, and carbon dioxide is a *nonpolar molecule.*

A water molecule (H_2O) is formed when two hydrogen atoms bind covalently to a single oxygen atom. Each hydrogen atom shares an electron pair with the oxygen atom, and again the oxygen has the stronger electron-attracting ability. But in this case, the molecule formed is V-shaped (H╲O╱H). The two hydrogen atoms are located at one end of the molecule, and the oxygen atom is at the other (Figure 2.8b); consequently, the electron pairs are not shared equally. This arrangement allows them to spend more time in the vicinity of the oxygen atom. Because electrons are negatively charged, that end of the molecule becomes slightly more negative (indicated by δ^-, the Greek letter delta with a minus sign) and the hydrogen end becomes slightly more positive (indicated by δ^+). In other words, a *polar molecule*, a molecule with two charged *poles*, is formed.

Polar molecules orient themselves toward other polar molecules or charged particles (ions, proteins, and others), and they play an important role in chemical reactions that occur in body cells. Because body tissues are 60 to 80 percent water, the fact that water is a polar molecule is particularly significant.

Hydrogen Bonds **Hydrogen bonds** are extremely weak bonds formed when a hydrogen atom bound to one "electron-hungry" nitrogen or oxygen atom is attracted by another such atom, and the hydrogen atom forms a "bridge" between them. Electrons are not involved in hydrogen bonds, as they are in ionic and covalent bonds. Hydrogen bonding is common between water

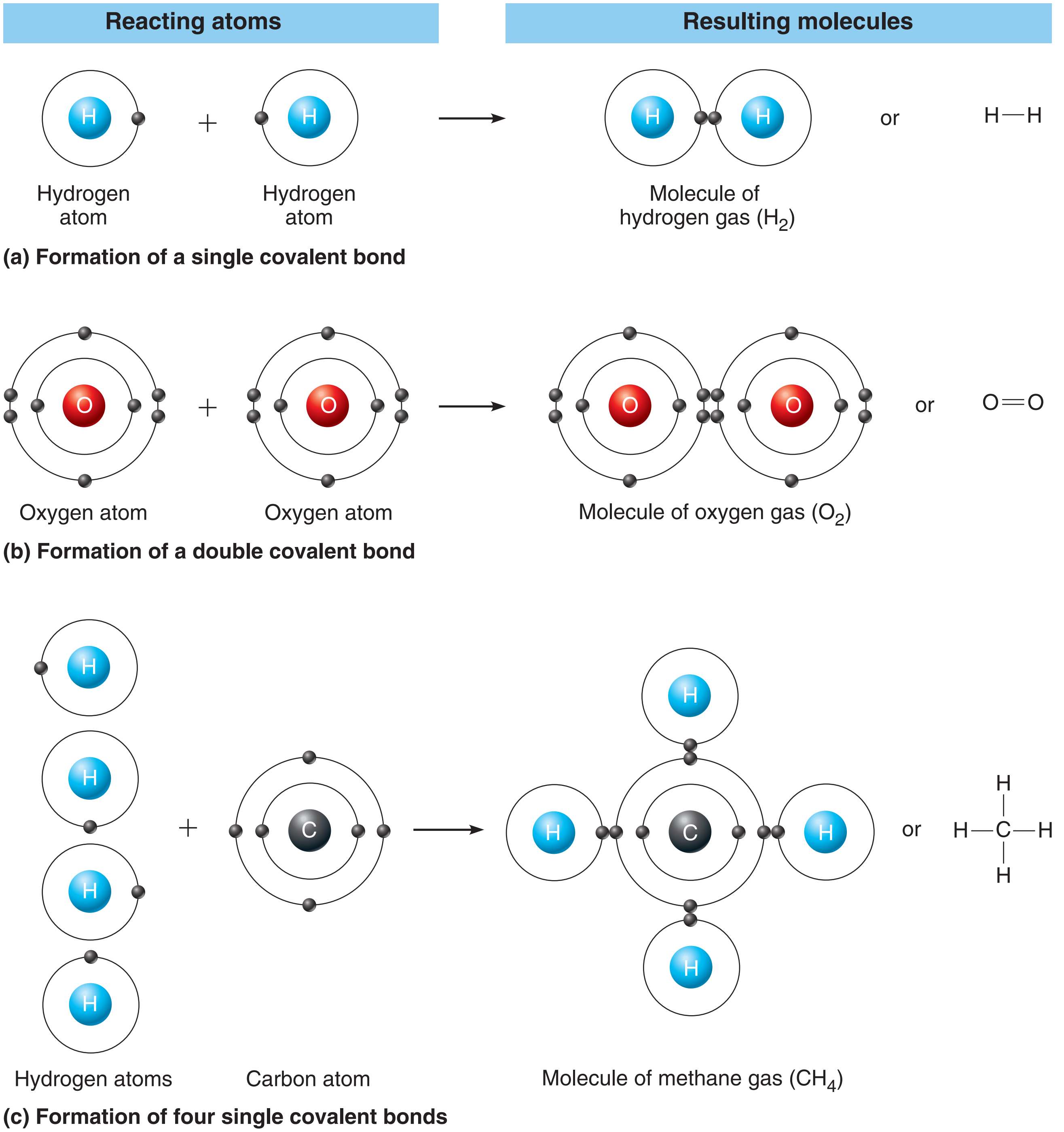
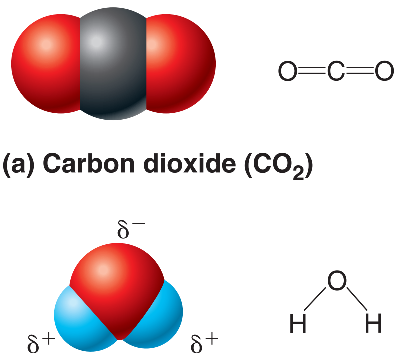


Figure 2.7 Formation of covalent bonds. (a) Formation of a single covalent bond between two hydrogen atoms forms a molecule of hydrogen gas. **(b)** Formation of a molecule of oxygen gas. Each oxygen atom shares two electron pairs with its partner; thus, a double covalent bond is formed. **(c)** Formation of a molecule of methane. A carbon atom shares four electron pairs with four hydrogen atoms. In the structural formulas of the molecules shown at the far right, each pair of shared electrons is indicated by a single line connecting the sharing atoms.

Q: ***Which molecule—(a) or (b)—is a polar molecule?***



(a) Carbon dioxide (CO_2)

(b) Water (H_2O)

Figure 2.8 Molecular models illustrating the three-dimensional structure of carbon dioxide and water molecules.

molecules **(Figure 2.9a)** and is reflected in water's surface tension. The surface tension of water causes it to "ball up," or form spheres at the surface. This tension allows some insects, such as water striders, to walk on water as long as they tread lightly (Figure 2.9b).

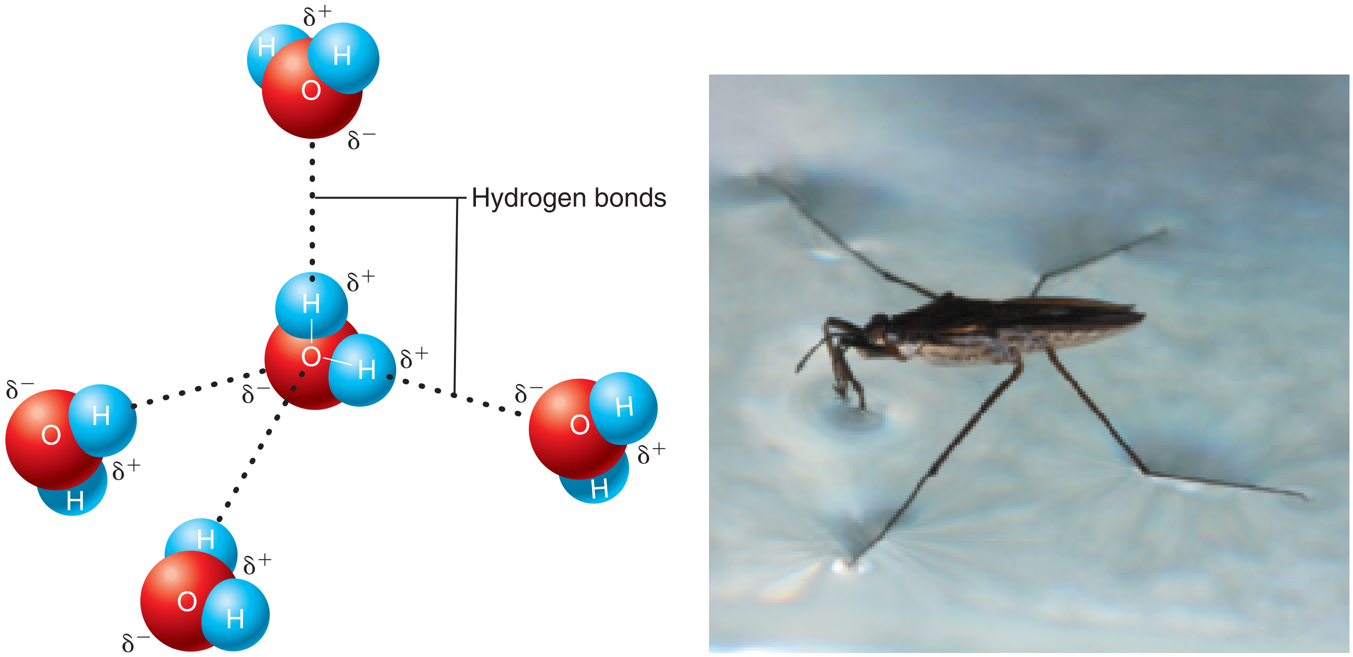
Hydrogen bonds are also important *intramolecular bonds*; that is, they help to bind different parts of the *same* molecule together into a special three-dimensional shape. These rather fragile bonds are very important in helping to maintain the structure of both protein molecules, which are essential functional molecules and body-building materials, and DNA, the genetic information molecule.

Patterns of Chemical Reactions

Chemical reactions involve the making or breaking of bonds between atoms. The total number of atoms remains the same, but the atoms appear in new combinations. Most chemical reactions have one of the three recognizable patterns we describe next.

Synthesis Reactions

Synthesis reactions occur when two or more atoms or molecules combine to form a larger,



(a)

(b)

Figure 2.9 Hydrogen bonding between polar water molecules. **(a)** The slightly positive ends of the water molecules (indicated by δ^+) become aligned with the slightly negative ends (indicated by δ^-) of other water molecules. **(b)** Water's high surface tension, a result of the combined strength of its hydrogen bonds, allows a water strider to walk on a pond without breaking the surface.

A: Water is a polar molecule.

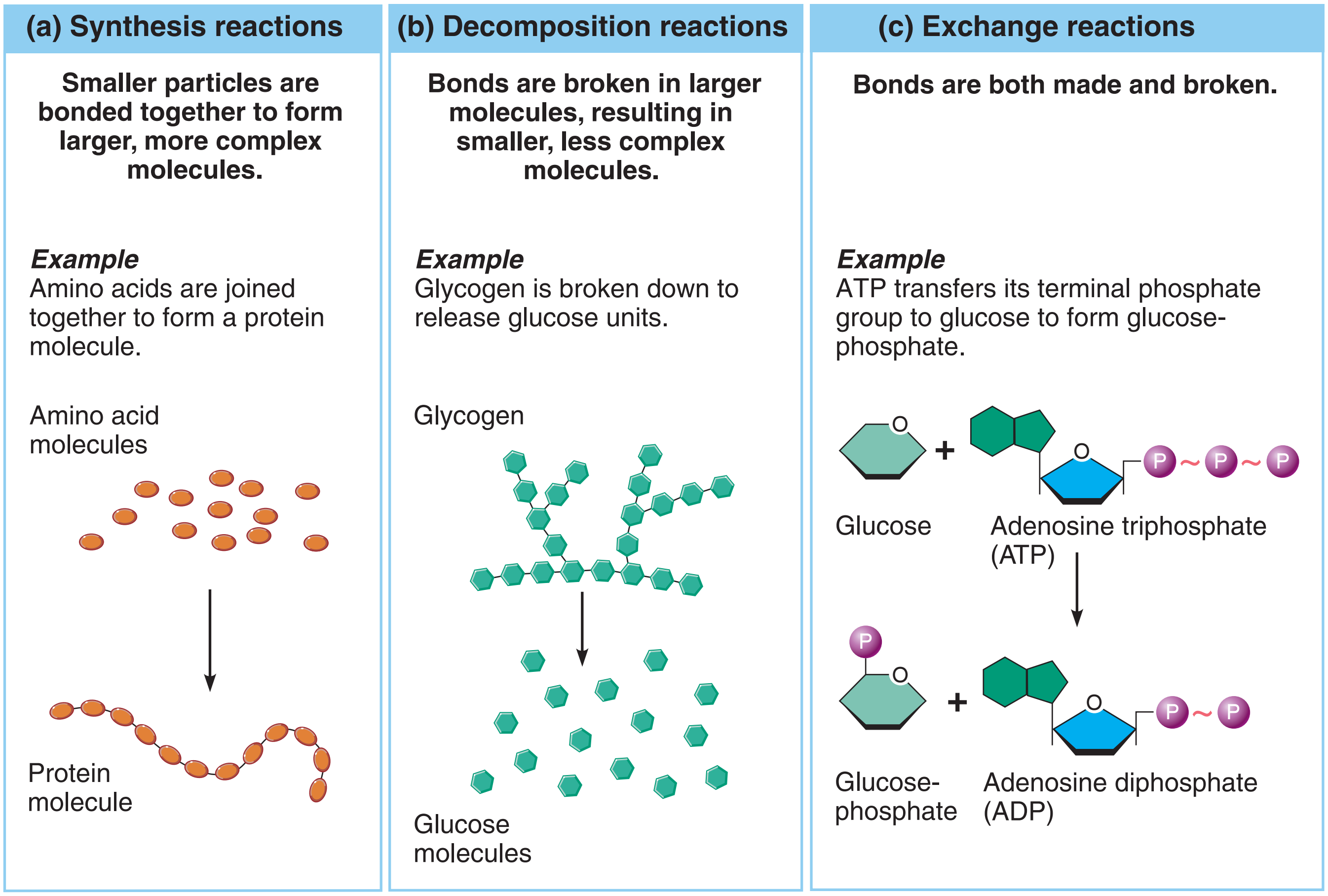


Figure 2.10 Patterns of chemical reactions.

more complex molecule, which can be simply represented as

$$A + B \rightarrow AB$$

Synthesis reactions always involve bond formation. Because energy must be absorbed to make bonds, synthesis reactions are energy-storing reactions.

Synthesis reactions underlie all anabolic (building) activities that occur in body cells. They are particularly important for growth and for repair of worn-out or damaged tissues. The formation of a protein molecule by the joining of *amino acids* (protein building blocks) into long chains is a synthesis reaction **(Figure 2.10a)**.

Decomposition Reactions

Decomposition reactions occur when a molecule is broken down into smaller molecules, atoms, or ions and can be indicated by

$$AB \rightarrow A + B$$

Essentially, decomposition reactions are synthesis reactions in reverse. Bonds are always broken, and the products of these reactions are smaller and simpler than the original molecules. As bonds are broken, chemical energy is released.

Decomposition reactions underlie all catabolic (destructive) processes that occur in body cells; that is, they are reactions that break down molecules. Examples include the digestion of foods into their building blocks and the breakdown of glycogen (a large carbohydrate molecule stored in the liver) to release glucose (Figure 2.10b) when the blood sugar level starts to decline.

Exchange Reactions

Exchange reactions involve simultaneous synthesis and decomposition reactions; in other words, bonds are both made and broken. During exchange reactions, a switch is made between molecule parts (changing partners, so to speak),

Table 2.4 Factors Increasing the Rate of Chemical Reactions

Factor	Mechanism to increase the number of collisions
↑ temperature	↑ the kinetic energy of the molecules, which in turn move more rapidly and collide more forcefully.
↑ concentration of reacting particles	↑ the number of collisions because of increased numbers of reacting particles.
↓ particle size	Smaller particles have more kinetic energy and move faster than larger ones, hence they take part in more collisions.
Presence of catalysts	↓ the amount of energy the molecules need to interact by holding the reactants in the proper positions to interact (see p. 51).

and different molecules are made. Thus, an exchange reaction can be generally indicated as

$$AB + C \rightarrow AC + B \quad \text{and} \quad AB + CD \rightarrow AD + CB$$

An exchange reaction occurs, for example, when ATP reacts with glucose and transfers its end phosphate group (its full name is *adenosine triphosphate*; *tri-* = three) to glucose, forming glucose-phosphate (Figure 2.10c). At the same time, the ATP becomes ADP, adenosine diphosphate (*di-* = two). This important reaction, which occurs whenever glucose enters a body cell, traps glucose inside the cell.

Most chemical reactions are reversible. If chemical bonds can be made, they can be broken, and vice versa. Reversibility is indicated by a double arrow in a chemical equation. When the arrows differ in length, the longer arrow indicates the more rapid reaction or the major direction in which the reaction is proceeding. For example, in the reaction

$$A + B \rightleftharpoons AB$$

the reaction going to the right is occurring more rapidly, so over time AB will accumulate while A and B will decrease in amount. If the arrows are of equal length, the reaction is at chemical equilibrium. Thus, in

$$A + B \rightleftharpoons AB$$

for each molecule of AB made, a molecule of AB is breaking down to release A and B.

Factors Influencing the Rate of Chemical Reactions

As mentioned earlier, for the atoms in molecules to react chemically, their outermost electron shells must overlap. In fact, for them to get close enough for this to happen, the particles must collide forcefully. Remember also that atoms move constantly because of their kinetic energy—this is what drives collisions between particles. Several factors, including temperature, concentration of the particles, and size of the particles, influence the kinetic energy and, hence, the speed of the particles and the force of collisions **(Table 2.4)**.

Did You Get It?

11. How do ionic bonds differ from covalent bonds?
12. What kind of bond forms between water molecules?
13. Which reaction type (see Figure 2.10) occurs when fats are digested in your small intestine?
14. How can you indicate that a chemical reaction is reversible?

For answers, see Appendix A.

Biochemistry: The Chemical Composition of Living Matter

→ **Learning Objective**

☐ **Distinguish organic from inorganic compounds.**

All chemicals found in the body fall into one of two major classes of molecules: either inorganic or organic compounds. The class of the compound is determined solely by the presence or absence of carbon. Except for a few so far unexplainable exceptions (such as carbon dioxide gas [CO_2] and carbon monoxide [CO]), **inorganic compounds** lack carbon and tend to be small, simple molecules. Examples of inorganic compounds found in the body are *water*, *salts*, and many (but not all) *acids* and *bases*. **Organic compounds** contain carbon. The important organic compounds in the body are *carbohydrates*, *lipids*, *proteins*, and *nucleic acids*. All organic compounds are fairly (or very) large covalent molecules.

Inorganic and organic compounds are equally essential for life. Trying to determine which is more valuable can be compared to trying to decide whether the ignition system or the engine is more essential to the operation of a car.

Inorganic Compounds

➔ Learning Objectives

- ☐ **Explain the importance of water to body homeostasis, and provide several examples of the roles of water.**
- ☐ **List several salts (or their ions) vitally important to body functioning.**
- ☐ **Differentiate a salt, an acid, and a base.**
- ☐ **Explain the concept of pH, and state the pH of blood.**

Water

Water is the most abundant inorganic compound in the body. It accounts for about two-thirds of body weight. Among the properties that make water so vital are the following:

- **High heat capacity.** Water has a *high heat capacity*; that is, it absorbs and releases large amounts of heat before its temperature changes. Thus, it prevents the sudden changes in body temperature that might otherwise result from intense sun exposure, chilling winter winds, or internal events (such as vigorous muscle activity) that liberate large amounts of heat.
- **Polarity/solvent properties.** Water is an excellent solvent because of its polarity; indeed, it is often called the "universal solvent." A *solvent* is a liquid or gas in which smaller amounts of other substances, called *solutes* (which may be gases, liquids, or solids), can be dissolved or suspended. The resulting mixture is called a *solution* when the solute particles are exceedingly tiny, and a *suspension* when the solute particles are fairly large. Translucent mixtures with solute particles of intermediate size are called *colloids*.

  Small reactive chemicals—such as salts, acids, and bases—dissolve easily in water and become evenly distributed. Molecules cannot react chemically unless they are in solution, so virtually all chemical reactions that occur in the body depend on water's solvent properties.

  Because nutrients, respiratory gases (oxygen and carbon dioxide), and wastes can dissolve in water, water can act as a transport and exchange medium in the body. For example, all these substances are carried around the body in blood plasma (the liquid part of blood that is mostly made up of water) and are exchanged between the blood and body cells by passing through the water-based interstitial fluid that bathes cells. Specialized molecules that lubricate the body, such as mucus and synovial fluid, also use water as their solvent. Synovial fluid "oils" the ends of bones as they move within joint cavities.
- **Chemical reactivity.** Water is an important *reactant* in some types of chemical reactions. For example, to digest foods or break down biological molecules, water molecules are added to the bonds of the larger molecules in order to break them. Such reactions are called *hydrolysis reactions*, a term that specifically recognizes this role of water (*hydro* = water; *lys* = splitting).
- **Cushioning.** Water also serves a protective function. In cerebrospinal fluid, water forms a cushion around the brain that helps to protect it from physical trauma. Amniotic fluid, which surrounds a developing fetus within the mother's body, plays a similar role in protecting the fetus.

Salts

A **salt** is an ionic compound containing cations other than the *hydrogen ion* (H^+) and anions other than the *hydroxide ion* (OH^-). Salts of many metal elements are commonly found in the body, but the most plentiful salts are those containing calcium and phosphorus, found chiefly in bones and teeth. When dissolved in body fluids, salts easily separate into their ions. This process, called *dissociation*, occurs rather easily because the ions have already been formed. All that remains is for the ions to "spread out." This is accomplished by the polar water molecules, which orient themselves with their slightly negative ends toward the cations and their slightly positive ends toward the anions, thereby overcoming the attraction between them **(Figure 2.11)**.

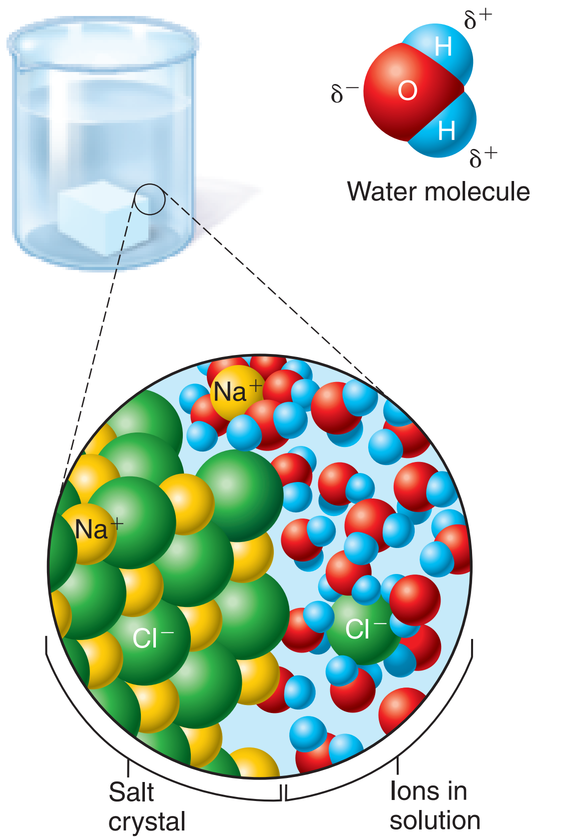


Figure 2.11 Dissociation of salt in water. The slightly negative ends of the water molecules (δ−) are attracted to Na^+, whereas the slightly positive ends of water molecules (δ+) orient toward Cl^-, causing the ions of the salt crystal to be pulled apart.

Salts, both in their ionic forms and in combination with other elements, are vital to body functioning. For example, sodium and potassium ions are essential for nerve impulses, and iron forms part of the hemoglobin molecule that transports oxygen within red blood cells.

Because ions are charged particles, all salts are **electrolytes**—substances that conduct an electrical current in solution. When electrolyte balance is severely disturbed, virtually nothing in the body works. (The functions of the elements found in body salts are summarized in Table 2.1 on p. 27.)

Acids and Bases

Like salts, acids and bases are electrolytes. That is, they ionize, dissociate in water, and can then conduct an electrical current.

Characteristics of Acids **Acids** have a sour taste and can dissolve many metals or "burn" a hole in your rug. Acids can have a devastating effect; for example, consider the damage to sea life, trees, and famous historical monuments caused by the vinegar-like acid rain. But the most useful definition of an acid is that it is a substance that can release *hydrogen ions (H^+)* in detectable amounts. Because a hydrogen ion is a hydrogen nucleus (a "naked proton"), acids are also defined as **proton (H^+) donors**. You may find it useful to think of acids as putting protons "in the game." As free protons, hydrogen ions can influence the acidity of body fluids.

When acids are dissolved in water, they release hydrogen ions and some anions. The anions are unimportant; it is the release of protons that determines an acid's effects on the environment. The ionization of hydrochloric acid (an acid produced by stomach cells that aids digestion) is shown in the following equation:

$$\underset{\text{(hydrochloric acid)}}{HCl} \rightarrow \underset{\text{(proton)}}{H^+} + \underset{\text{(anion)}}{Cl^-}$$

Other acids found or produced in the body include acetic acid (the acidic component of vinegar) and carbonic acid.

Acids that ionize completely and liberate all their protons are called *strong acids*; an example is hydrochloric acid. Acids that ionize incompletely, as do acetic and carbonic acid, are called *weak acids*. For example, when carbonic acid dissolves in water, only some of its molecules ionize to liberate H^+.

$$\underset{\text{(carbonic acid)}}{2H_2CO_3} \rightarrow \underset{\text{(proton)}}{H^+} + \underset{\text{(anion)}}{HCO_3^-} + \underset{\text{(carbonic acid)}}{H_2CO_3}$$

Characteristics of Bases **Bases** have a bitter taste, feel slippery, and are **proton (H^+) acceptors**. (You can also think of them as taking protons "out of the game"; when protons are bound to a molecule, they are unable to affect the acidity of body fluids.) Hydroxides are common inorganic bases. Like acids, the hydroxides ionize and dissociate in water; but in this case, the *hydroxide ion (OH^-)* and some cations are released. The ionization of sodium hydroxide (NaOH), commonly known as lye, is shown as

$$\underset{\text{(sodium hydroxide)}}{NaOH} \rightarrow \underset{\text{(cation)}}{Na^+} + \underset{\text{(hydroxide ion)}}{OH^-}$$

The hydroxide ion is an avid proton (H^+) seeker, and any base containing this ion is considered a strong base. By contrast, *bicarbonate ion* (HCO_3^-), an important base in blood, is a fairly weak base.

Acids, Bases, and Neutralization When acids and bases are mixed, they react with each other (in an exchange reaction) to form water and a salt:

$$\underset{\text{(acid)}}{HCl} + \underset{\text{(base)}}{NaOH} \rightarrow \underset{\text{(water)}}{H_2O} + \underset{\text{(salt)}}{NaCl}$$

This type of exchange reaction, in which an acid and a base interact, is more specifically called a **neutralization** reaction.

pH: Acid-Base Concentrations The relative concentration of hydrogen (and hydroxide) ions in various body fluids is measured in concentration units called **pH** (pe-āch) units. The **pH scale**, which was devised in 1909 by a Danish biochemist (and part-time beer brewer) named Sørensen, is based on the number of protons in solution.

The pH scale runs from 0 to 14 **(Figure 2.12)**, and each successive change of 1 pH unit represents a tenfold change in hydrogen ion concentration. At a pH of 7, the number of hydrogen ions exactly equals the number of hydroxide ions, and the solution is neutral; that is, neither acidic nor basic. Solutions with a pH lower than 7 are acidic: The hydrogen ions outnumber the hydroxide ions. A solution with a pH of 6 has 10 times as many hydrogen ions as a solution with a pH of 7, and a pH of 3 indicates a 10,000-fold (10 × 10 × 10 × 10) increase in hydrogen ion concentration from pH 7. Solutions with a pH number higher than 7 are basic, or alkaline, and solutions with a pH of 8 and 12 have 1/10 and 1/100,000 (respectively) the number of hydrogen ions present in a solution with a pH of 7.

Q: ***Which ion is responsible for increased acidity?***

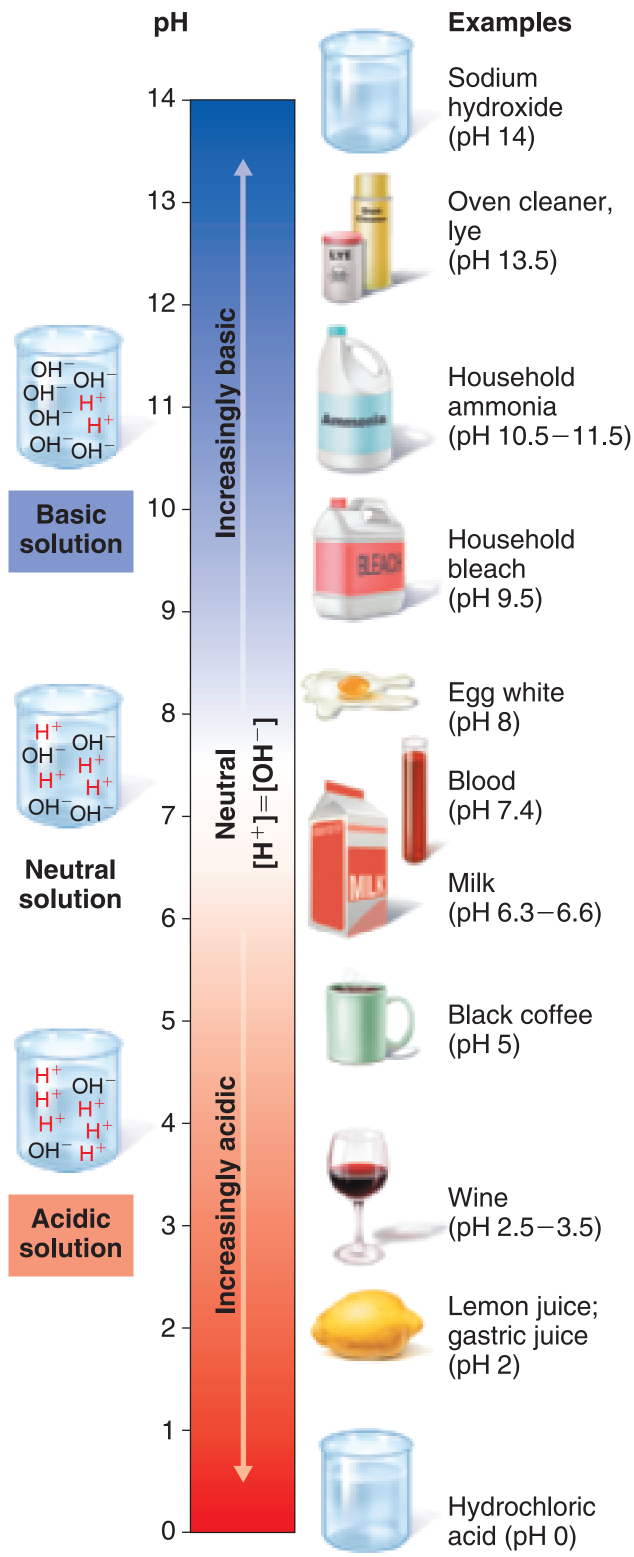


Figure 2.12 The pH scale and pH values of representative substances. The pH scale is based on the number of hydrogen ions in solution. At a pH of 7, the number of H^+ = the number of OH^-, and the solution is neutral. A solution with a pH below 7 is acidic (more H^+ than OH^-); above 7, basic, or alkaline (less H^+ than OH^-).

A: The hydrogen ion.

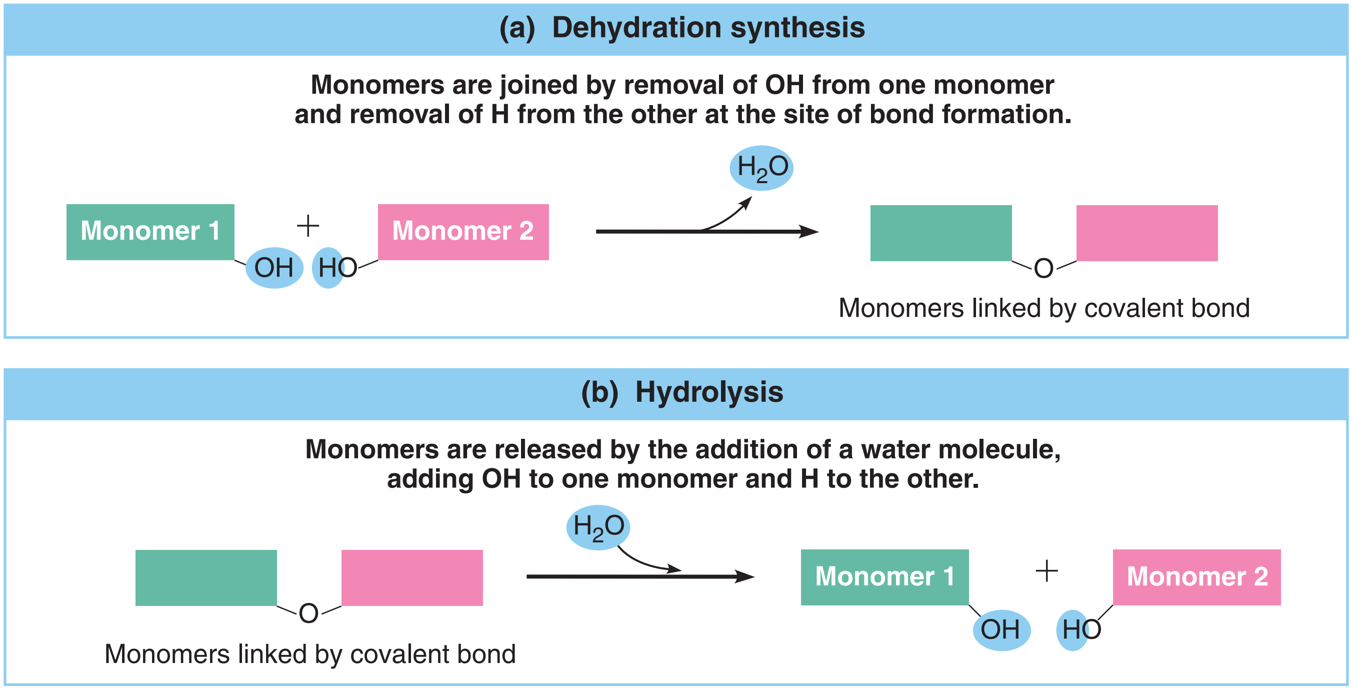


Figure 2.13 Dehydration synthesis and hydrolysis of biological molecules. Biological molecules are formed from their monomers (units) by dehydration synthesis and broken down to their monomers by hydrolysis.

Living cells are extraordinarily sensitive to even slight changes in pH. Acid-base balance is carefully regulated by the kidneys, lungs, and a number of chemicals called **buffers** that are present in body fluids. Weak acids and weak bases are important components of the body's buffer systems, which act to maintain pH stability by taking up excess hydrogen or hydroxide ions (see Chapter 15).

Because blood comes into close contact with nearly every body cell, regulation of blood pH is especially critical. Normally, blood pH varies in a narrow range, from 7.35 to 7.45. When blood pH changes more than a few tenths of a pH unit from these limits, death becomes a distinct possibility. Although we could give hundreds of examples to illustrate this point, we will provide just one very important one: When blood pH begins to dip into the acid range, the amount of life-sustaining oxygen that the hemoglobin in blood can carry to body cells decreases rapidly to dangerously low levels.

Did You Get It?

15. What property of water prevents rapid changes in body temperature?
16. Which is a proton donor—an acid or a base?
17. Is a pH of 11 acidic or basic? What is the difference in pH between a solution at pH 11 and a solution at pH 5?
18. Biochemistry is "wet" chemistry. What does this statement mean?
19. Salts are electrolytes. What does that mean?

For answers, see Appendix A.

Organic Compounds

➔ Learning Objectives

- ☐ **Explain the role of dehydration synthesis and hydrolysis in formation and breakdown of organic molecules.**
- ☐ **Compare and contrast carbohydrates and lipids in terms of their building blocks, structures, and functions in the body.**

Most organic compounds are very large molecules, but their interactions with other molecules typically involve only small, reactive parts of their structure called *functional groups* (acid groups, amines, and others).

Many organic compounds (carbohydrates and proteins, for example) are polymers. **Polymers** are chainlike molecules made of many similar or repeating units (**monomers**), which are joined together by **dehydration synthesis** (Figure 2.13a). During dehydration synthesis, a hydrogen atom is

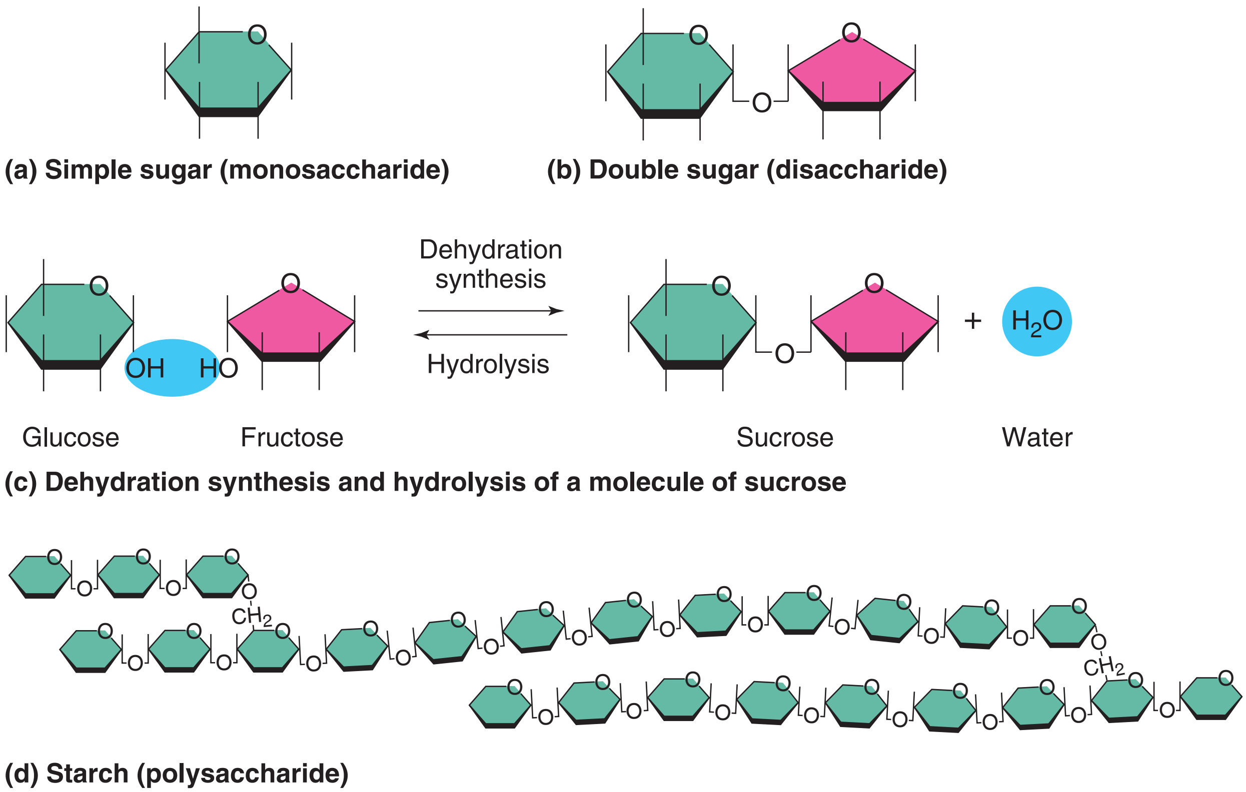


Figure 2.14 Carbohydrates. (a) The generalized structure of a monosaccharide. **(b)** and **(d)** The basic structures of a disaccharide and a polysaccharide, respectively. **(c)** Formation and breakdown of the disaccharide sucrose by dehydration synthesis and hydrolysis, respectively.

removed from one monomer and a hydroxyl group (OH) is removed from the monomer it is to be joined with. A covalent bond then forms, uniting the monomers, and a water molecule is released. This removal of a water molecule (dehydration) at the bond site occurs each time a monomer is added to the growing polymer chain.

When polymers must be broken down or digested to their monomers, the reverse process, called **hydrolysis**, occurs (Figure 2.13b). As a water molecule is added to each bond, the bond is broken, releasing the monomers. All organic molecules covered in this chapter—carbohydrates, lipids, proteins, and nucleic acids—are formed by dehydration synthesis and broken down by hydrolysis.

Carbohydrates

Carbohydrates, which include sugars and starches, contain carbon, hydrogen, and oxygen. With slight variations, the hydrogen and oxygen atoms appear in the same ratio as in water; that is, two hydrogen atoms to one oxygen atom. This is reflected in the word *carbohydrate*, which means "hydrated carbon," and in the molecular formulas of sugars. For example, glucose is $C_6H_{12}O_6$, and ribose is $C_5H_{10}O_5$.

Carbohydrates are classified according to size and solubility in water as monosaccharides, disaccharides, or polysaccharides. Because monosaccharides are joined to form the molecules of the other two groups, they are the structural units, or building blocks, of carbohydrates.

Monosaccharides **Monosaccharide** means one (*mono*) sugar (*saccharide*), and thus monosaccharides are also referred to as *simple sugars*. They are single-chain or single-ring structures (meaning the carbon backbone forms either a line or a circle), containing from three to seven carbon atoms (Figure 2.14a).

The most important monosaccharides in the body are glucose, fructose, galactose, ribose, and deoxyribose. **Glucose**, also called *blood sugar*, is the universal cellular fuel. *Fructose* and *galactose* are converted to glucose for use by body cells.

Ribose and *deoxyribose* form part of the structure of nucleic acids, another group of organic molecules responsible for genetic information.

Disaccharides **Disaccharides**, or *double sugars* (Figure 2.14b), are formed when two simple sugars are joined by dehydration synthesis. In this reaction, as noted earlier, a water molecule is lost as the bond forms (Figure 2.14c).

Some of the important disaccharides in the diet are *sucrose* (glucose-fructose), which is cane sugar; *lactose* (glucose-galactose), found in milk; and *maltose* (glucose-glucose), or malt sugar. Because the double sugars are too large to pass through cell membranes, they must be broken down (digested) to their monosaccharide units to be absorbed from the digestive tract into the blood; this is accomplished by hydrolysis (see Figures 2.14b and 2.13b).

Polysaccharides Long, branching chains of linked simple sugars are called **polysaccharides** (literally, "many sugars") (Figure 2.14d). Because they are large, insoluble molecules, they are ideal storage products. Another consequence of their large size is that they lack the sweetness of simple and double sugars.

Only two polysaccharides, starch and glycogen, are of major importance to the body. *Starch* is the storage polysaccharide formed by plants. We ingest it in the form of "starchy" foods, such as grain products (corn, rice) and root vegetables (potatoes and carrots, for example). *Glycogen* is a slightly smaller, but similar, polysaccharide found in animal tissues (largely in the muscles and the liver). Like starch, it is a polymer of linked glucose units.

Carbohydrates provide a ready, easily used source of fuel for cells, and glucose is at the top of the "cellular menu." When glucose is oxidized ("burned" by combining with oxygen) in a complex set of chemical reactions, it is broken down into carbon dioxide and water. Some of the energy released as the glucose bonds are broken is trapped in the bonds of high-energy ATP molecules, the energy "currency" of all body cells. If not immediately needed for ATP synthesis, dietary carbohydrates are converted to glycogen or fat and stored. Those of us who have gained weight from eating too many carbohydrate-rich snacks have firsthand experience of this conversion process!

Small amounts of carbohydrates are used for structural purposes and represent 1 to 2 percent of cell mass. Some sugars are found in our genes, and others are attached to outer surfaces of cell membranes (boundaries), where they act as road signs to guide cellular interactions.

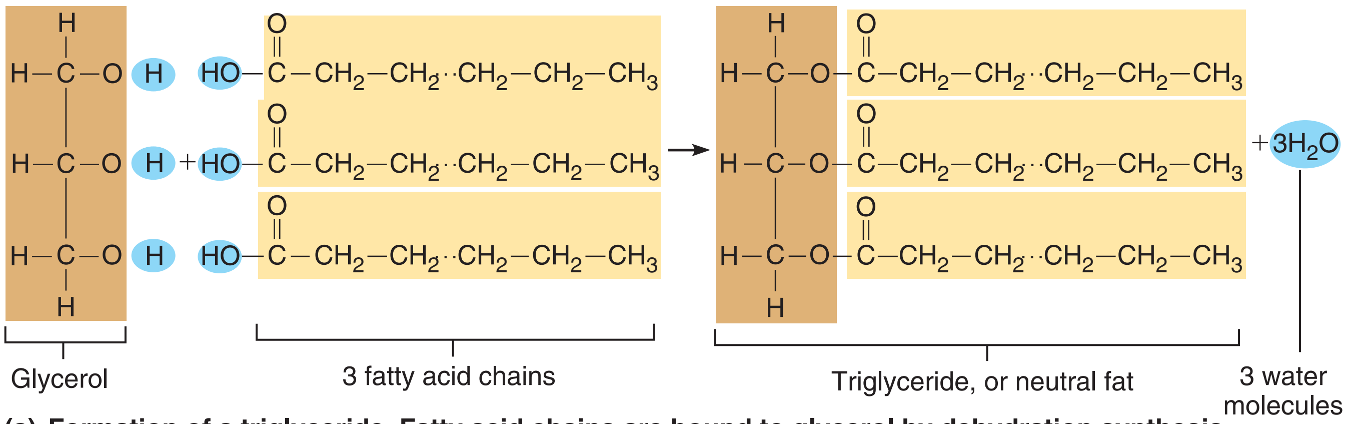
Lipids

Lipids, or fats, are a large and diverse group of organic compounds (**Table 2.5**, p. 47). They enter the body in the form of meats, egg yolks, dairy products, and oils. The most abundant lipids in the body are *triglycerides*, *phospholipids*, and *steroids*. Like carbohydrates, all lipids contain carbon, hydrogen, and oxygen atoms, but in lipids, carbon and hydrogen atoms far outnumber oxygen atoms, as illustrated by the formula for a typical fat named tristearin, which is $C_{57}H_{110}O_6$. Lipids are insoluble in water but readily dissolve in other lipids and in organic solvents such as alcohol and acetone.

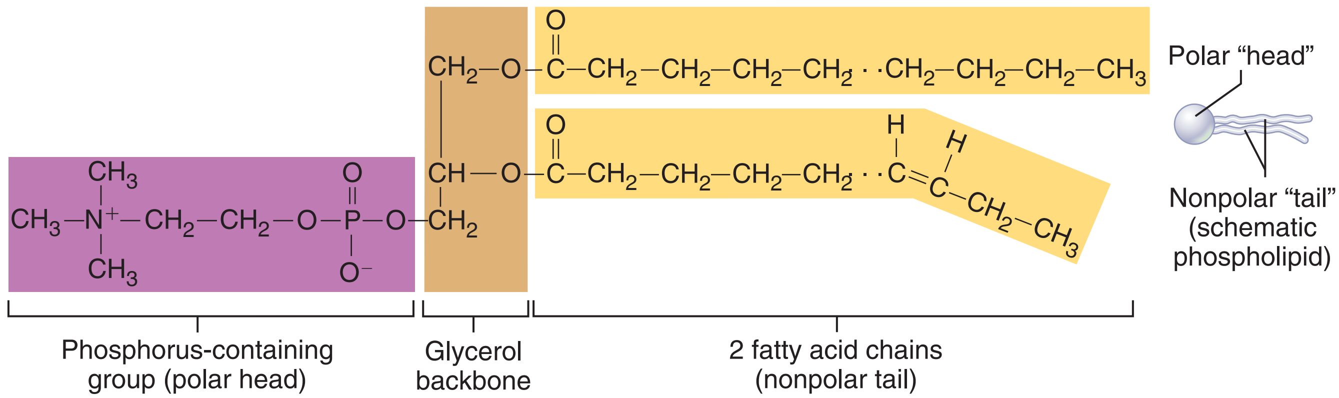
Triglycerides The **triglycerides** (tri-glis′er-īdz), or **neutral fats**, are composed of two types of building blocks, **fatty acids** and **glycerol**. Their synthesis involves the attachment of three fatty acids to a single glycerol molecule. The result is an E-shaped molecule that resembles the tines of a fork **(Figure 2.15a)**. Although the glycerol backbone is the same in all neutral fats, the fatty acid chains vary; this variation results in different kinds of neutral fats.

The length of a triglyceride's fatty acid chains and their type of C—C bonds determine how solid the molecule is at a given temperature. Fatty acid chains with only single covalent bonds between carbon atoms are referred to as **saturated fats**. Their fatty acid chains are straight (**Figure 2.16a**, p. 46), and, at room temperature, the molecules of a saturated fat pack closely together, forming a solid; butter is an example. Fatty acids that contain one or more double bonds between carbon atoms are said to be **unsaturated fats** (**monounsaturated** and **polyunsaturated**, respectively). The double and triple bonds cause the fatty acid chains to kink (Figure 2.16b) so that they cannot pack closely enough to solidify. Hence, triglycerides with short fatty acid chains or unsaturated fatty acids are oils (liquid at room temperature). Plant lipids are typically oils. Examples include olive oil (rich in monounsaturated fats) and soybean and safflower oils, which contain a high percentage of

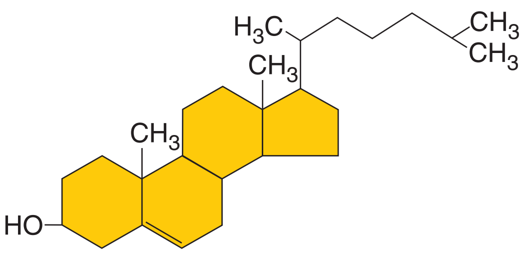
Q: ***Triglycerides and phospholipids are similar. What is the major structural difference between them?***



(a) Formation of a triglyceride. Fatty acid chains are bound to glycerol by dehydration synthesis.



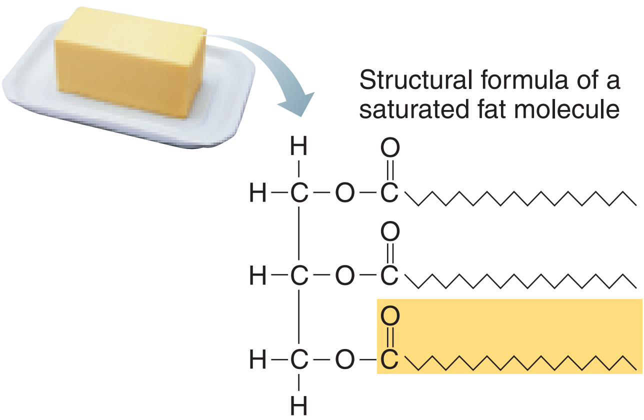
(b) Typical structure of a phospholipid molecule (phosphatidylcholine). Two fatty acid chains and a phosphorous-containing group are attached to a glycerol backbone.



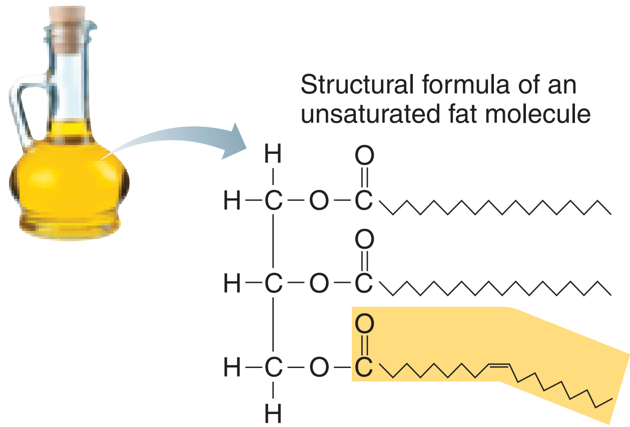
(c) Cholesterol. Simplified structure of cholesterol, a steroid, formed by four interlocking carbon rings.

Figure 2.15 Lipids. (a) Triglycerides, or neutral fats, are synthesized by dehydration synthesis, and a water molecule is lost at each bond site. **(b)** Structure of a typical phospholipid molecule. Notice that one fatty acid chain in (b) is unsaturated (has one or more C═C [double] bonds). **(c)** The generalized structure of cholesterol (the basis for all steroids made in the body).

A: They both have a glycerol backbone and fatty acid chains. However, triglycerides have three attached fatty acid chains, whereas phospholipids have only two; the third is replaced by a negatively charged phosphate-containing group.



(a) Saturated fat. At room temperature, the molecules of a saturated fat such as this butter are packed closely together, forming a solid.



(b) Unsaturated fat. At room temperature, the molecules of an unsaturated fat such as this olive oil cannot pack together closely enough to solidify because of the kinks in some of their fatty acid chains.

Figure 2.16 Examples of saturated and unsaturated fats and fatty acids.

polyunsaturated fatty acids. Longer fatty acid chains and more saturated fatty acids are common in animal fats, such as butterfat and the fat of meats, which are solid at room temperature. Of the two types of fatty acids, the unsaturated variety, especially olive oil, is more "heart healthy."

Trans fats, common in many margarines and baked products, are oils that have been solidified by the addition of hydrogen atoms at sites of double carbon bonds, which reduces those bonds to single carbon bonds. They increase the risk of heart disease by raising "bad" cholesterol and decreasing "good" cholesterol. In contrast, the **omega-3 fatty acids**, found naturally in cold-water fish, appear to decrease the risk of heart disease and some inflammatory diseases.

Triglycerides represent the body's most abundant and concentrated source of usable energy. When they are oxidized, they yield large amounts of energy. They are stored chiefly in fat deposits beneath the skin and around body organs, where they help insulate the body and protect deeper body tissues from heat loss and injury.

Phospholipids. **Phospholipids** (fos′fo-lip″idz) are similar to triglycerides. The major difference is that a phosphorus-containing group is always part of the molecule and takes the place of one of the fatty acid chains. Thus, phospholipids have two instead of three attached fatty acids (see Figure 2.15b).

Because the phosphorus-containing portion (the "head") bears a negative charge, phospholipids have special chemical properties and polarity. For example, the charged region is **hydrophilic** ("water loving"), meaning it attracts and interacts with water and ions. The fatty acid chains (the nonpolar "tail") are **hydrophobic** ("water fearing") and do not interact with polar or charged molecules. The presence of phospholipids in *cell membranes* allows cells to be selective about what may enter or leave.

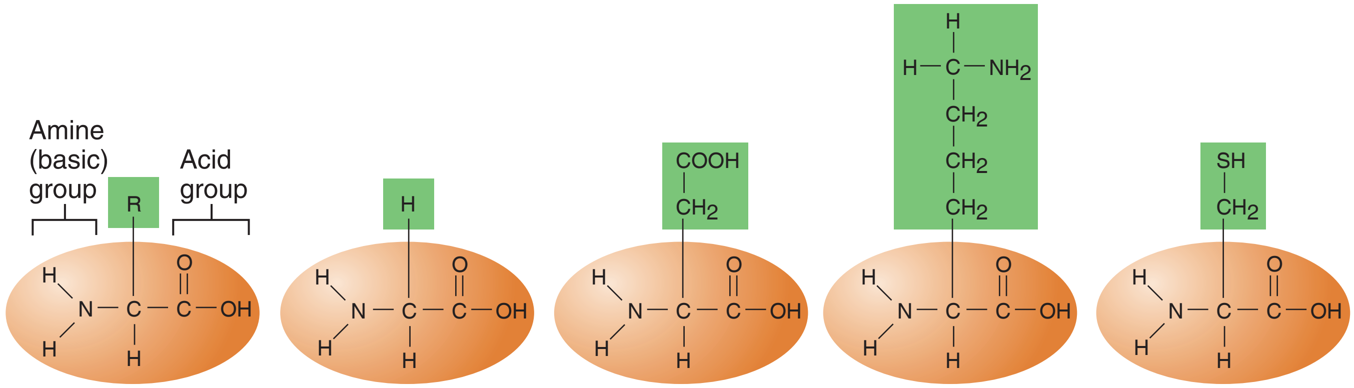
Steroids **Steroids** are basically flat molecules formed of four interlocking carbon rings (see Figure 2.15c); thus, their structure differs quite a bit from that of other lipids. However, like fats, steroids are made largely of hydrogen and carbon atoms and are fat-soluble. The single most important steroid molecule is **cholesterol**. We ingest cholesterol in animal products such as meat, eggs, and cheese, and some is made by the liver, regardless of dietary intake.

Cholesterol has earned bad press because of its role in arteriosclerosis, but it is essential for human life. Cholesterol is found in cell membranes and is the raw material used to make vitamin D, steroid hormones (chemical signaling molecules in the body), and bile salts (which help break down fats during digestion). Although steroid hormones are present in the body in only small quantities, they are vital to homeostasis. Without sex hormones, reproduction would be impossible; and without the corticosteroids produced by the adrenal glands, the body would not survive.

Table 2.5 Representative Lipids Found in the Body

Lipid type	Location/function
Triglycerides (neutral fats)	Found in fat deposits (subcutaneous tissue and around organs); protect and insulate the body organs; the major source of stored energy in the body.
Phospholipids	Found in cell membranes; participate in the transport of lipids in plasma; abundant in the brain and the nervous tissue in general, where they help to form insulating white matter.
Steroids	
Cholesterol	The basis of all body steroids.
Bile salts	A breakdown product of cholesterol; released by the liver into the digestive tract, where they aid in fat digestion and absorption.
Vitamin D	A fat-soluble vitamin produced in the skin on exposure to UV (ultraviolet) radiation (sunshine); necessary for normal bone growth and function.
Sex hormones	Estrogen and progesterone (female hormones) and testosterone (a male hormone) produced from cholesterol; necessary for normal reproductive function; deficits result in sterility.
Corticosteroids (adrenal cortical hormones)	Cortisol, a glucocorticoid, is a long-term antistress hormone that is necessary for life; aldosterone helps regulate salt and water balance in body fluids by targeting the kidneys.
Other lipid-based substances	
Fat-soluble vitamins	
A	Found in orange-pigmented vegetables (carrots) and fruits (tomatoes); part of the photoreceptor pigment involved in vision.
E	Taken in via plant products such as wheat germ and green leafy vegetables; may promote wound healing and contribute to fertility, but not proven in humans; an antioxidant; may help to neutralize free radicals (highly reactive particles believed to be involved in triggering some types of cancers).
K	Made available largely by the action of intestinal bacteria; also prevalent in a wide variety of foods; necessary for proper clotting of blood.
Prostaglandins	Derivatives of fatty acids found in cell membranes; various functions depending on the specific class, including stimulation of uterine contractions (thus inducing labor and miscarriages), regulation of blood pressure, and control of motility of the gastrointestinal tract; involved in inflammation.
Lipoproteins	Lipoid and protein-based substances that transport fatty acids and cholesterol in the bloodstream; major varieties are high-density lipoproteins (HDLs) and low-density lipoproteins (LDLs).
Glycolipids	Component of cell membranes. Lipids associated with carbohydrate molecules that determine blood type, play a role in cell recognition or in recognition of foreign substances by immune cells.

Q: ***What is the importance of the R-groups (green here)?***



(a) Generalized structure of all amino acids.

(b) Glycine is the simplest amino acid.

(c) Aspartic acid (an acidic amino acid) has an acid group (—COOH) in the R group.

(d) Lysine (a basic amino acid) has an amine group (—NH_2) in the R group.

(e) Cysteine (a basic amino acid) has a sulfhydryl (—SH) group in the R group.

Figure 2.17 Amino acid structures. (a) Generalized structure of amino acids. All amino acids have both an amine (NH_2) group and an acid (COOH) group; they differ only in the atomic makeup of their R-groups (green). **(b–e)** Specific structures of four amino acids.

Did You Get It?

20. What are the structural units, or building blocks, of carbohydrates? Of lipids?
21. Which type(s) of lipid(s) is/are found in cellular membranes?
22. Is the formation of glycogen from glucose an example of hydrolysis or dehydration synthesis?

For answers, see Appendix A.

Proteins

→ **Learning Objectives**

- ☐ **Differentiate fibrous proteins from globular proteins.**
- ☐ **Define *enzyme*, and explain the role of enzymes.**

Proteins account for over 50 percent of the organic matter in the body, and they have the most varied functions of all organic molecules. Some are construction materials; others play vital roles in cell function. Like carbohydrates and lipids, all proteins contain carbon, oxygen, and hydrogen. In addition, they contain nitrogen and sometimes sulfur atoms as well.

The building blocks of proteins are small molecules called **amino** (ah-me′no) **acids**. Twenty varieties of amino acids are found in human proteins. All amino acids have an *amine group* (NH_2), which gives them basic properties, and an *acid group* (COOH), which allows them to act as acids. In fact, all amino acids are identical except for a single group of atoms called their *R-group* (Figure 2.17). Differences in the R-groups make each amino acid chemically unique. For example, an extra acid group (COOH) in the R-group (such as in aspartic acid) makes the amino acid more acidic, and a sulfhydryl group (—SH) in the R-group (such as in cysteine) is often involved in stabilizing protein structure by binding to another amino acid within the same protein.

Amino acids are joined together in chains to form *polypeptides* (fewer than 50 amino acids); proteins (more than 50 amino acids); and large, complex proteins (50 to thousands of amino acids). Because each type of amino acid has distinct properties, the sequence in which they are

A: The R-groups make each type of amino acid functionally unique.

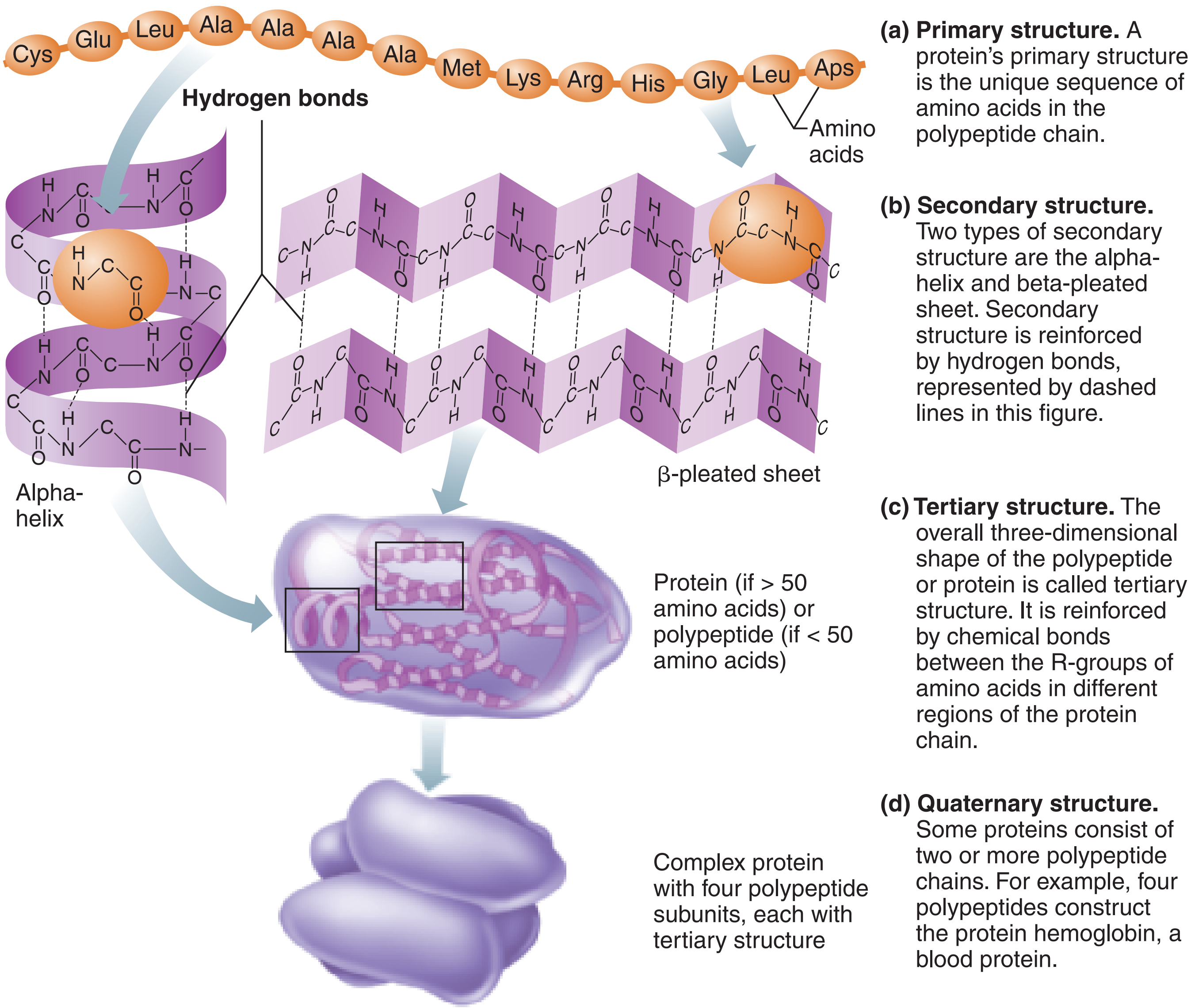
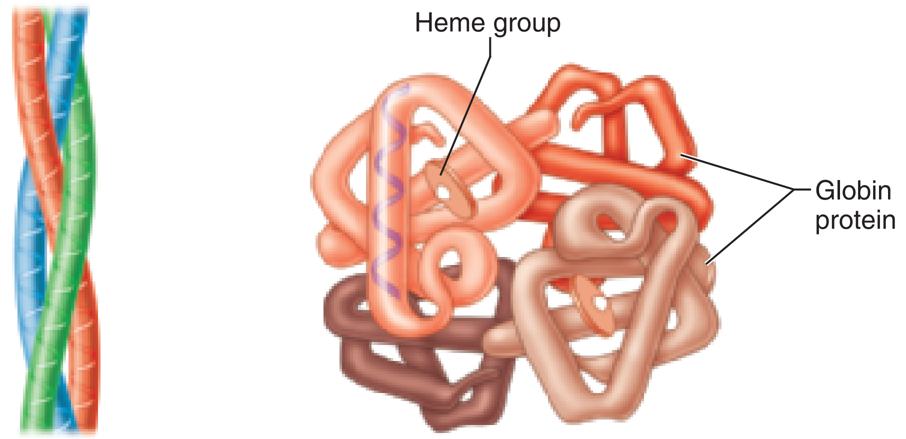


Figure 2.18 **The four levels of protein structure.**

bound together produces proteins that vary widely both in structure and function. To understand this more easily, think of the 20 amino acids as a 20-letter alphabet. The letters (amino acids) are used in specific combinations to form words (a protein). Just as a change in one letter of any word can produce a word with an entirely different meaning (flour → floor) or one that is nonsensical (flour → fluur), changes in amino acids (letters) or in their positions in the protein allow literally thousands of different protein molecules to be made.

Structural Levels of Proteins Proteins can be described in terms of four structural levels. The sequence of amino acids composing each amino acid chain is called the *primary structure*. This structure, which resembles a strand of amino acid "beads," is the backbone of a protein molecule in which the chemical properties of each amino acid will affect how the protein folds **(Figure 2.18a)**.

Most proteins do not function as simple, linear chains of amino acids. Instead, they twist or bend upon themselves to form a more complex *secondary structure*. The most common secondary structure is the **alpha (α)-helix**, which resembles a metal spring (Figure 2.18b). The α-helix is formed by coiling of the primary chain and is stabilized by hydrogen bonds. Hydrogen bonds in α-helices



(a) Triple helix of collagen (a fibrous or structural protein).

(b) Hemoglobin molecule composed of the protein globin and attached heme groups. (Globin is a globular, or functional, protein.)

Figure 2.19 General structure of (a) a fibrous protein and (b) a globular protein.

always link different parts of the *same* chain together.

In another type of secondary structure, the **beta (β)-pleated sheet**, the primary polypeptide chains do not coil, but are linked side by side by hydrogen bonds to form a pleated, ribbonlike structure that resembles the pleats of a skirt or a sheet of paper folded into a fan (see Figure 2.18b). In this type of secondary structure, the hydrogen bonds may link together *different polypeptide chains* as well as *different parts* of the same chain that has folded back on itself.

Many proteins have *tertiary structure* (ter′she-a″re), the next higher level of complexity. Tertiary structure is achieved when α-helical or β-pleated regions of the amino acid chain fold upon one another to produce a compact ball-like, or *globular*, protein (Figure 2.18c). The unique structure is maintained by covalent and hydrogen bonds between amino acids that are often far apart in the primary chain. Finally, when two or more amino acid chains (polypeptide chains) combine in a regular manner to form a complex protein, the protein has *quaternary* (kwah′ter-na″re) *structure* (Figure 2.18d).

Although a protein with tertiary or quaternary structure looks a bit like a crumpled ball of tin foil, the final structure of any protein is very specific and is dictated by its primary structure. In other words, the types and positions of amino acids in the protein backbone determine where hydrogen bonds can form to keep water-loving (hydrophilic) amino acids near the surface and water-fearing (hydrophobic) amino acids buried in the protein's core so the protein remains water-soluble.

Fibrous and Globular Proteins Based on their overall shape and structure, proteins are classed as either fibrous or globular proteins **(Figure 2.19)**. The strandlike **fibrous proteins**, also called **structural proteins**, appear most often in body structures. Some exhibit only secondary structure, but most have tertiary or even quaternary structure. They are very important in binding structures together and providing strength in certain body tissues. For example, *collagen* (kol′ah-jen) is found in bones, cartilage, and tendons and is the most abundant protein in the body (Figure 2.19a). *Keratin* (ker′ah-tin) is the structural protein of hair and nails and the material that makes skin tough.

Globular proteins are mobile, generally compact, spherical molecules that have at least tertiary structure. These water-soluble proteins play crucial roles in virtually all biological processes. Because they *do things* rather than just form structures, they are also called **functional proteins**; the scope of their activities is remarkable **(Table 2.6)**. For example, some proteins called antibodies help to provide immunity; others (hormones) help to regulate growth and development. Still others, called

Table 2.6 **Representative Classes of Functional Proteins**

Functional class	Role(s) in the body
Antibodies (immunoglobulins)	Highly specialized proteins that recognize, bind with, and inactivate bacteria, toxins, and some viruses; function in the immune response, which helps protect the body from "invading" foreign substances.
Hormones	Help to regulate growth and development. Examples include • Growth hormone—an anabolic hormone necessary for optimal growth. • Insulin—helps regulate blood sugar levels. • Nerve growth factor—guides the growth of neurons in the development of the nervous system.
Transport proteins	Hemoglobin transports oxygen in the blood; other transport proteins in the blood carry iron, cholesterol, or other substances.
Enzymes (catalysts)	Essential to virtually every biochemical reaction in the body; increase the rates of chemical reactions by at least a millionfold; in their absence (or destruction), biochemical reactions cease.

enzymes (en′zīmz), regulate essentially every chemical reaction that goes on within the body. The oxygen-carrying protein hemoglobin is an example of a globular protein with quaternary structure (Figure 2.19b).

The fibrous structural proteins are exceptionally stable, but the globular functional proteins are quite the opposite. Hydrogen bonds are critically important in maintaining their structure, but hydrogen bonds are fragile and are easily broken by heat and excesses of pH. When their three-dimensional structures are destroyed, the proteins are said to be *denatured* and can no longer perform their physiological roles. Why? Their function depends on their specific three-dimensional shapes. Hemoglobin becomes totally unable to bind and transport oxygen when blood pH becomes too acidic, as we stated earlier. Pepsin, a protein-digesting enzyme that acts in the stomach, is inactivated by alkaline pH. In each case, the improper pH has destroyed the structure required for function.

Because enzymes are important in the functioning of all body cells, we consider these incredibly complex molecules here.

Enzymes and Enzyme Activity **Enzymes** are functional proteins that act as biological catalysts. A **catalyst** is a substance that increases the rate of a chemical reaction without becoming part of the product or being changed itself. Enzymes can accomplish this feat because they have unique regions called **active sites** on their surfaces. These sites "fit" and interact chemically with other molecules of complementary shape and charge called *substrates* **(Figure 2.20)**. While the substrates are bound to the enzyme's active site, producing a structure called an *enzyme-substrate complex*, they undergo structural changes that result in a new product. Whereas some enzymes build larger molecules, others break things into smaller pieces or simply modify a substrate. Once the reaction has occurred, the enzyme releases the product. Because enzymes are not changed during the reaction, they are reusable, and the cells need only small amounts of each enzyme. Think of scissors cutting paper. The scissors (the enzyme) are unchanged while the paper (the substrate) is "cut" to become the product. The scissors are reusable and go on to "cut" other paper.

Enzymes are capable of catalyzing millions of reactions each minute. However, they do more than just increase the speed of chemical reactions; they also determine just which reactions are possible at a particular time. No enzyme, no reaction! Enzymes can be compared to a bellows used to fan a sluggish fire into flaming activity. Without enzymes, biochemical reactions would occur far too slowly to sustain life.

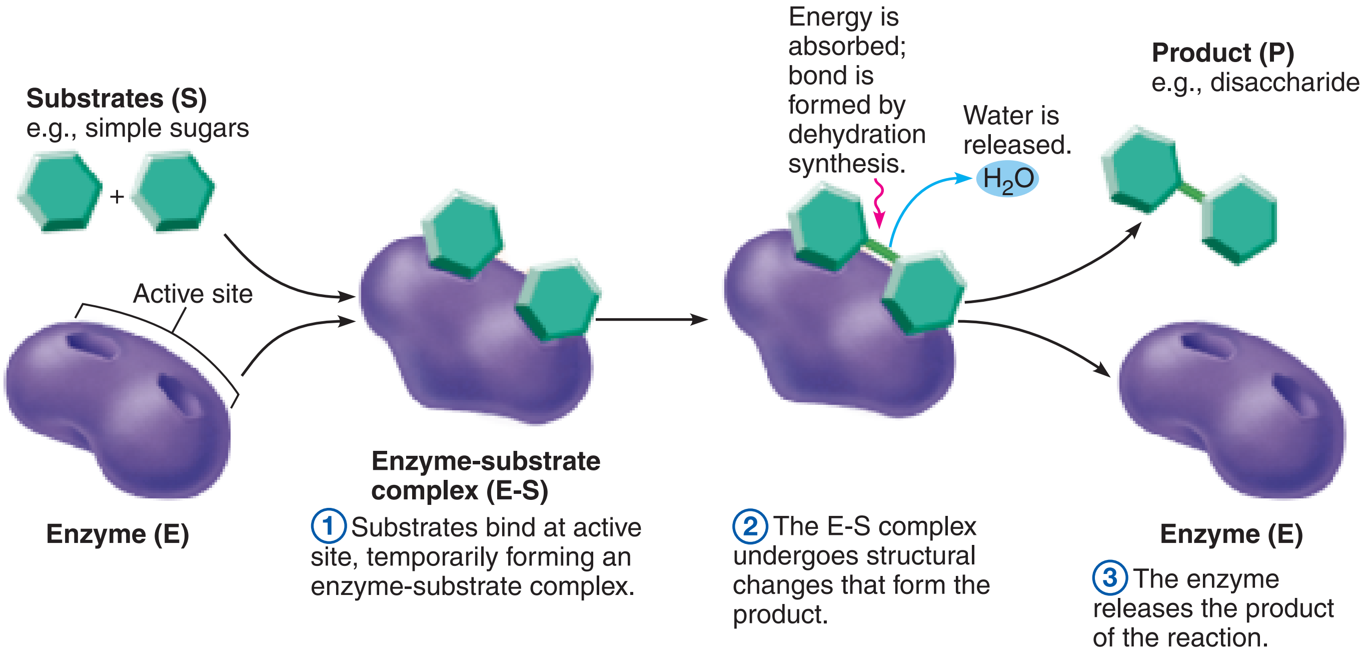


Figure 2.20 A simplified view of enzyme action.

Although there are hundreds of different kinds of enzymes in body cells, they are very specific in their activities, each controlling only one (or a small group of) chemical reaction(s) and acting only on specific molecules. Most enzymes are named according to the specific type of reaction they catalyze. For example, *hydrolases* add water, and *oxidases* cause oxidation. (In most cases, you can recognize an enzyme by the suffix **-ase** in its name.)

The activity of different enzymes is controlled in different ways. Many enzymes are produced in an inactive form and must be activated in some way before they can function. In other cases, enzymes are inactivated immediately after they have performed their catalytic function. Both events are true of enzymes that promote blood clotting when a blood vessel has been damaged. If this were not so, large numbers of unneeded and potentially lethal blood clots would be formed.

Did You Get It?

23. What is the primary structure of proteins?
24. Which is more important for building body structures, fibrous or globular proteins?
25. How does an enzyme recognize its substrate(s)?

For answers, see Appendix A.

Nucleic Acids

➔ Learning Objectives

- ☐ **Compare and contrast the structures and functions of DNA and RNA.**
- ☐ **Explain the importance of ATP in the body.**

The role of **nucleic** (nu-kle′ik) **acids** is fundamental: They make up your genes, which provide the basic blueprint of life. They not only determined what type of organism you would be, but also directed your growth and development—and they did this largely by dictating protein structure. (Remember that enzymes, which catalyze all the chemical reactions that occur in the body, are proteins.)

Nucleic acids, composed of carbon, oxygen, hydrogen, nitrogen, and phosphorus atoms, are the largest biological molecules in the body. Their building blocks, **nucleotides** (nu′kle-o-tīdz), are quite complex. Each consists of three basic parts: (1) a nitrogen-containing base, (2) a pentose (5-carbon) sugar, and (3) a phosphate group (**Figure 2.21a** and **b**).

The bases come in five varieties: *adenine* (A), *guanine* (G), *cytosine* (C), *thymine* (T), and *uracil* (U). A and G are large, nitrogen-containing bases made up of two carbon rings, whereas C, T and U are smaller, single-ring structures. The nucleotides

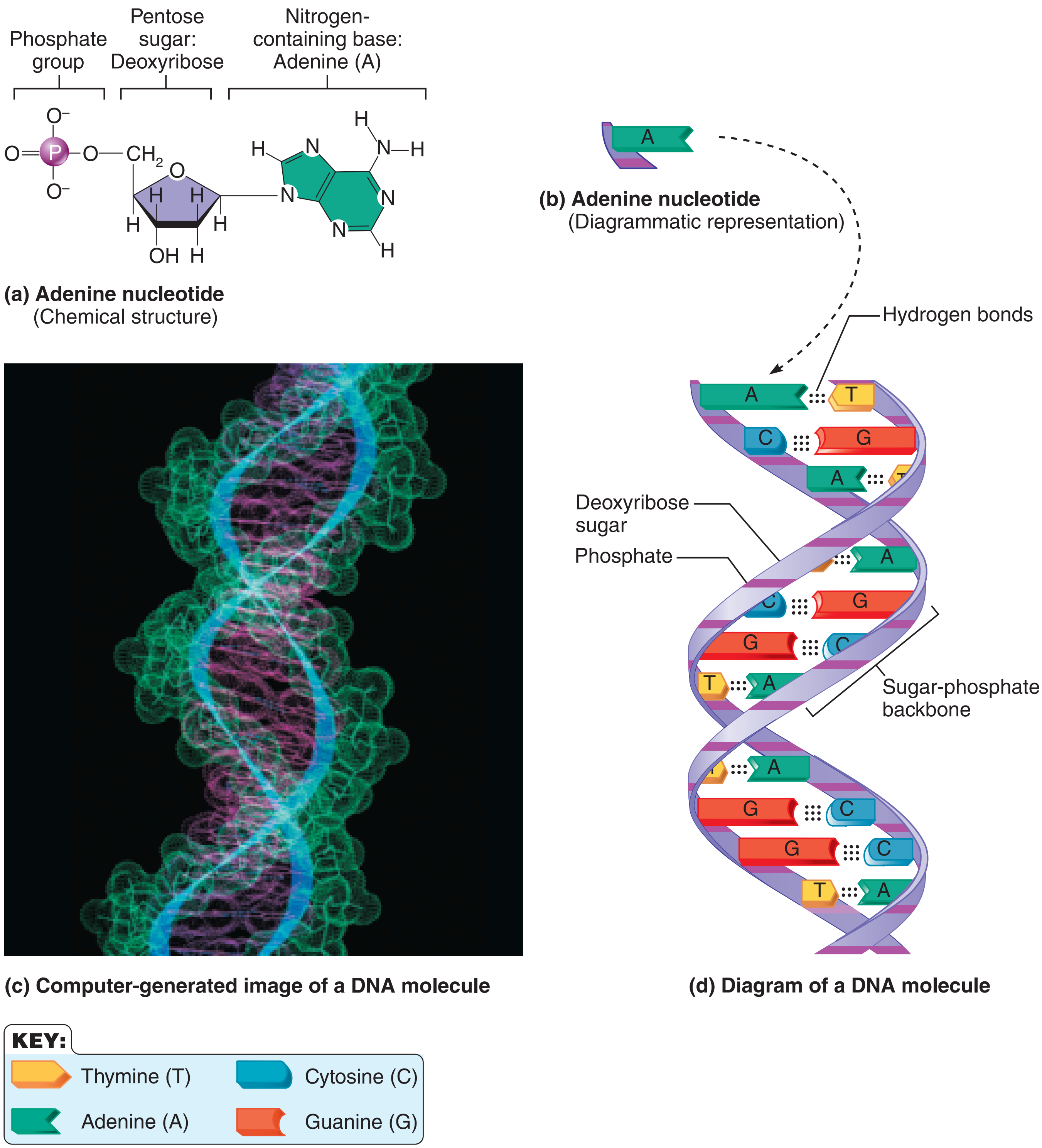
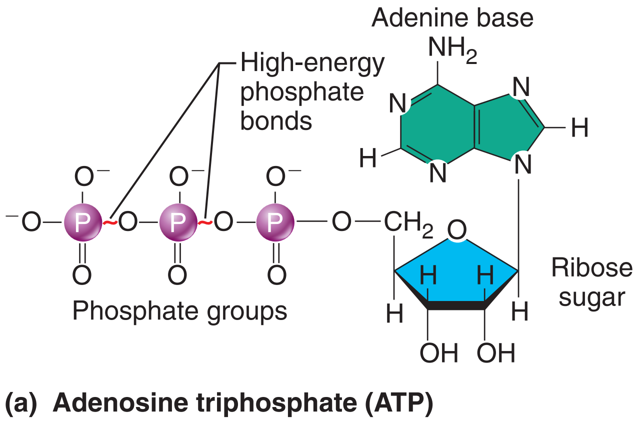


Figure 2.21 Structure of DNA. **(a)** The unit of DNA, the nucleotide, is composed of a deoxyribose sugar molecule linked to a phosphate group. A nitrogen-containing base is attached to the sugar. The nucleotide illustrated, both in its **(a)** chemical and **(b)** diagrammatic structures, contains the base adenine. **(c)** Computer-generated image of DNA. **(d)** Diagrammatic structure of a DNA molecule—two nucleotide chains coiled into a double helix. The "backbones" of DNA are formed by alternating sugar and phosphate molecules. The "rungs" are formed by complementary bases (A to T, G to C) bound by two or three hydrogen bonds, respectively.



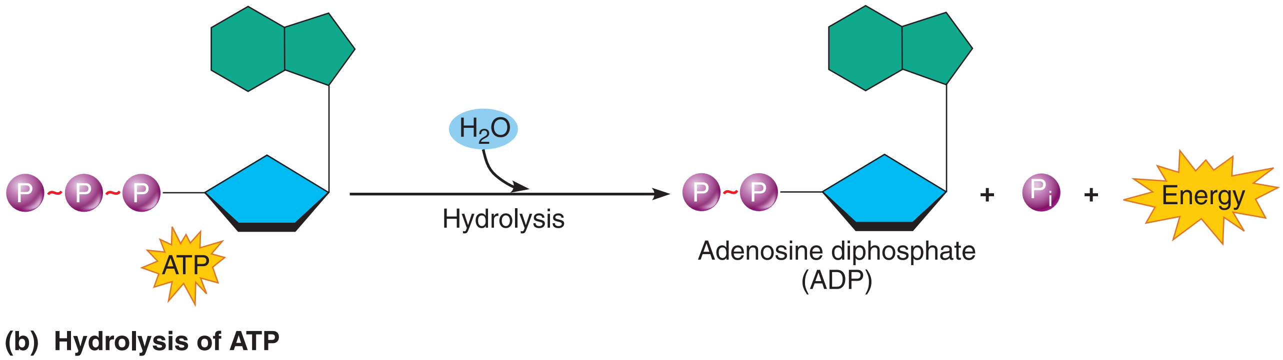


Figure 2.22 ATP—structure and hydrolysis. **(a)** The structure of ATP (adenosine triphosphate). **(b)** Hydrolysis of ATP to yield ADP (adenosine diphosphate) and inorganic phosphate (P_i). High-energy bonds are indicated by a red ~.

are named according to the base they contain: A-containing bases are adenine nucleotides, C-containing bases are cytosine nucleotides, and so on.

The two major kinds of nucleic acids are **deoxyribonucleic** (de-ok″sĭ-ri″bo-nu-kle′ik) **acid (DNA)** and **ribonucleic acid (RNA)**. DNA and RNA differ in many respects. DNA is the genetic material found within the cell nucleus (the control center of the cell). It has two fundamental roles: (1) It replicates itself exactly before a cell can divide, thus ensuring that every body cell gets an identical copy of the genetic information; and (2) it provides the instructions for building every protein in the body. For the most part, RNA functions outside the nucleus and can be considered the "molecular assistant" of DNA; that is, RNA carries out the orders for protein synthesis issued by DNA.

Although both DNA and RNA are formed when nucleotides join together, their final structures are different. DNA is a long double chain of nucleotides (Figure 2.21c and d). Its bases are A, G, T, and C, and its sugar is *deoxyribose*. Its two nucleotide chains are held together by hydrogen bonds between the bases, forming a ladderlike molecule. Alternating sugar and phosphate molecules form the "uprights" of the ladder, called the *sugar-phosphate backbone*, and each "rung" is formed by two joined bases (one *base pair*). Binding between the bases is very specific: A always binds to T, and G always binds to C. Thus, A and T are said to be *complementary bases*, as are C and G. Therefore, a sequence of ACTGA on one nucleotide chain would be bound to the complementary sequence TGACT on the other nucleotide strand. The whole molecule is then coiled into a spiral-staircase-like structure called a *double helix*.

Whereas DNA is double-stranded, RNA molecules are single nucleotide strands. The RNA bases are A, G, C, and U (U replaces the T found in DNA), and its sugar is *ribose* instead of deoxyribose. Three major varieties of RNA exist—*messenger*, *transfer*, and *ribosomal RNA*—and each has a specific role to play in carrying out DNA's instructions for building proteins. Messenger RNA (mRNA) carries the information for building the protein from the DNA to the ribosomes, the protein-synthesizing sites. Transfer RNA (tRNA) ferries amino acids to the ribosomes. Ribosomal RNA (rRNA) forms part of the ribosomes, where it oversees the translation of the message and the binding together of amino acids to form the proteins. (We describe protein synthesis in greater detail in Chapter 3.)

Adenosine Triphosphate (ATP)

The synthesis of **adenosine triphosphate** (ah-den′o-sēn tri-fos′fāt), or **ATP**, is all-important because it provides a form of chemical energy that all body cells can use. Without ATP, molecules cannot be made or broken down, cells cannot maintain their membrane boundaries, and all life processes grind to a halt.

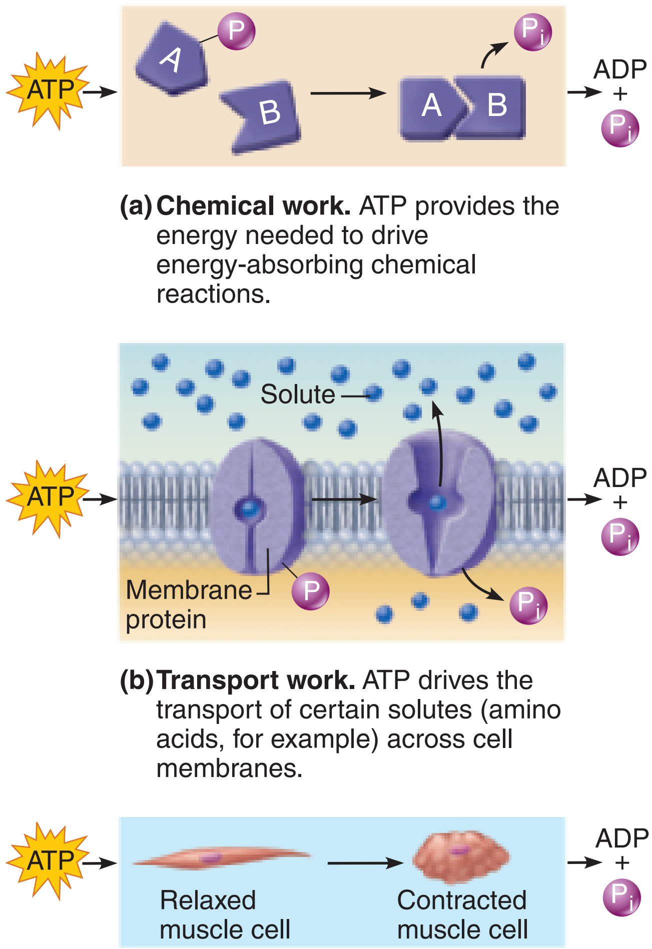
Although glucose is the most important fuel for body cells, none of the chemical energy contained in its bonds can be used directly to power cellular work. Instead, energy is released as glucose is oxidized, and it is captured and stored in the bonds of ATP molecules as small packets of energy. Compare glucose and ATP to crude oil and gasoline: While oil stores lots of energy, it cannot be used directly to power a car without first being refined into gasoline.

Structurally, ATP is a modified RNA nucleotide; it consists of an adenine base, ribose sugar, and three phosphate groups instead of one **(Figure 2.22a)**. The phosphate groups are attached by unique chemical bonds called *high-energy phosphate bonds*. These bonds are high energy because the phosphate groups are negatively charged and the like charges repel one another, generating tension at these sites. As ATP is used to provide cellular energy, **adenosine diphosphate (ADP)** accumulates (Figure 2.22b), and ATP supplies are replenished by oxidation of food fuels. Essentially, the same amount of energy must be captured and used to reattach a phosphate group to ADP (that is, to reverse the reaction) as is liberated when the terminal phosphate is removed from ATP.

When the high-energy bonds of ATP are broken by hydrolysis, energy is liberated and can be used immediately by the cell to do work or power a particular activity—such as driving chemical reactions, transporting solutes across the membrane, or, in the case of muscle cells, contracting **(Figure 2.23)**. ATP can be compared to a tightly coiled spring that is ready to uncoil with tremendous energy when the "catch" is released. The consequence of breaking its terminal phosphate bond can be represented as follows:

$$\text{ATP} \rightleftharpoons \text{ADP} + \text{P}_i + \text{E}$$

ATP		ADP		P_i		E
(adenosine triphosphate)		(adenosine diphosphate)		(inorganic phosphate)		(energy)



(a) Chemical work. ATP provides the energy needed to drive energy-absorbing chemical reactions.

(b) Transport work. ATP drives the transport of certain solutes (amino acids, for example) across cell membranes.

(c) Mechanical work. ATP activates contractile proteins in muscle cells so that the cells can shorten and perform mechanical work.

Figure 2.23 Three examples of how ATP drives cellular work. The high-energy bonds of ATP release energy for use by the cell when they are broken. ATP is regenerated (phosphate is again bound to ADP) as energy is released by the oxidation of food fuels and captured in the third high-energy bond.

Did You Get It?

26. How do DNA and RNA differ from each other in the kinds of bases and sugars they contain?
27. What is the vital importance of ATP to body cells?

For answers, see Appendix A.

FOCUS ON CAREERS

Pharmacy Technician

To recognize how medications affect patients, pharmacy technicians need a thorough understanding of anatomy and physiology.

When most people get a new medication, they open up the package and toss out the little pamphlet that goes into detail about how the medication works. Not Chris Green. "I love reading the package inserts," says Green, the lead pharmacy technician at a CVS drugstore in Birmingham, Alabama. Green's enthusiasm for those details is a lifesaver for his customers. Pharmacy technicians are a vital link in the chain between doctor and patient.

Although pharmacy technicians are legally prohibited from talking with patients about their symptoms, they can translate medical jargon and discuss a medication's side effects and other precautions the patient may need to take. For example, doctors may recommend that patients who are on certain medications for a long time have regular tests such as eye exams, bloodwork, or tests for liver function. A pharmacy technician can convey that information to the patient—and check on subsequent visits to make sure he or she is following up.

A busy retail pharmacy has various stations: data entry, where the patient's record is updated with a new prescription; production, where the prescription is filled; and verification, where the pharmacist reviews the prescription and makes sure it is filled and labeled correctly. Green's job is to make sure the process flows smoothly from station to station.

Pharmacy technicians must have a good grasp of anatomy and physiology to understand each drug's chemical properties.

Green started working as a cashier at a drugstore when he was in high school and gradually became interested in the pharmacy itself.

"I was interested in how drugs work, how they can help people and improve their health," he says.

Having earned a bachelor's degree in biology, Green emphasizes that pharmacy technicians must have a good grasp of the sciences, especially basic chemistry and anatomy and physiology, to help them understand each drug's chemical makeup and properties.

"When all the data is entered, we see what potential side effects there are," he says. "It's important to know how the medications work, how they interact with each other, how they interact with the body. I might see something and bring it to the pharmacist's attention." In addition, communication skills and the ability to work with people are important. Good communication can be the difference between life and death for a patient, particularly when a doctor prescribes a

medication that could react badly with another medication the patient is already taking. Drug interactions happen commonly when you have multiple doctors. "Sometimes, we'll get two ACE inhibitors in the same category from two different doctors [prescribed for the same patient], and that could be lethal," Green says.

Pharmacy technicians work in retail and mail-order pharmacies, hospitals, nursing homes, assisted living facilities, and anywhere else patients have high needs for medication. As the Baby Boom generation ages and the number of senior citizens grows, so does the demand for pharmacists and pharmacy technicians.

Requirements to be a pharmacy technician vary from state to state, and many aspiring technicians simply receive on-the-job training. However, some pharmacies seek out technicians with specific training requiring classroom and laboratory work in a hospital, community college, or vocational program or sometimes through the military. Some of these programs also include internships in pharmacies.

For additional information on this career and others, click the Focus on Careers link at MasteringA&P®.

Summary

Concepts of Matter and Energy (pp. 24–26)

1. Matter
   a. Matter is anything that occupies space and has mass.
   b. Matter exists in three states: gas, liquid, and solid.
2. Energy
   a. Energy is the capacity to do work or to move matter. Energy has kinetic (active) and potential (stored) work capacities.
   b. Energy forms that are important in body function include chemical, electrical, mechanical, and radiant.
   c. Energy forms can be converted from one form to another, but some energy is always unusable (lost as heat) in such transformations.

Composition of Matter (pp. 26–31)

1. Elements and atoms
   a. Each element is a unique substance that cannot be decomposed into simpler substances by ordinary chemical methods. A total of 118 elements exists; they differ from one another in their chemical and physical properties.
   b. Four elements (carbon, hydrogen, oxygen, and nitrogen) make up 96 percent of living matter. Several other elements are present in small or trace amounts.
   c. The building blocks of elements are atoms. Each atom is designated by an atomic symbol consisting of one or two letters.
2. Atomic structure
   a. Atoms are composed of three subatomic particles: protons, electrons, and neutrons. Protons are positively charged, electrons are negatively charged, and neutrons are neutral.
   b. The planetary model of the atom portrays all the mass of the atom (protons and neutrons) concentrated in a central nucleus. Electrons orbit the nucleus along specific orbits. The orbital model also locates protons and neutrons in a central nucleus, but it depicts electrons as occupying areas of space called orbitals and forming an electron cloud of negative charge around the nucleus.
   c. Each atom can be identified by an atomic number, which is equal to the number of protons contained in the atom's nucleus. Because all atoms are electrically neutral, the number of protons in any atom is equal to its number of electrons.
   d. The atomic mass number is equal to the sum of the protons and neutrons in the atom's nucleus.
   e. Isotopes are different atomic forms of the same element; they differ only in the number of neutrons. Many of the heavier isotopes are unstable and decompose to a more stable form by ejecting particles or energy from the nucleus, a phenomenon called radioactivity. Such radioisotopes are useful in medical diagnosis and treatment and in biochemical research.
   f. The atomic weight is approximately equal to the mass number of the most abundant isotope of any element.

Molecules and Compounds (pp. 31–32)

1. A molecule is the smallest unit resulting from the bonding of two or more atoms. If the atoms are different, a molecule of a compound is formed.
2. Compounds exhibit properties different from those of the atoms that comprise them.

Chemical Bonds and Chemical Reactions (pp. 32–38)

1. Bond formation
   a. Chemical bonds are energy relationships. Electrons in the outermost energy level (valence shell) of the reacting atoms are active in the bonding.
   b. Atoms with a stable valence shell (two electrons in shell 1, or eight in the subsequent shells) are chemically inactive. Those with an incomplete valence shell interact by losing, gaining, or sharing electrons to achieve stability (that is, to either fill the valence shell or meet the rule of eight).
   c. Ions are formed when valence electrons are completely transferred from one atom to another. The oppositely charged ions thus formed attract each other, forming an ionic bond. Ionic bonds are common in salts.
   d. Covalent bonds involve the sharing of electron pairs between atoms. If the electrons are shared equally, the molecule is a nonpolar covalent molecule. If the electrons are not shared equally, the molecule is a polar covalent molecule. Polar molecules orient themselves toward charged particles and other molecules.

e. Hydrogen bonds are fragile bonds that bind together water molecules or different parts of the same molecule (intramolecular bonds) but do not involve electrons. They are common in large, complex organic molecules, such as proteins and nucleic acids.

2. Patterns of chemical reactions

a. Chemical reactions involve the formation or breaking of chemical bonds. They are indicated by a chemical equation, which provides information about the atomic composition (formula) of the reactant(s) and product(s).

b. Chemical reactions that result in larger, more complex molecules are synthesis reactions; they involve storing energy in the bonds formed.

c. In decomposition reactions, larger molecules are broken down into simpler molecules or atoms. Bonds are broken, releasing energy.

d. Exchange reactions involve both the making and breaking of bonds. Atoms are replaced by other atoms.

e. Regardless of the type of reaction, most chemical reactions are reversible. Reversibility is indicated by a double arrow.

3. Factors increasing the rate of chemical reactions

a. For atoms to interact chemically, they must collide forcefully.

b. Factors that affect the number or force of collisions include the temperature, concentration of the reactants, particle size, and catalysts (enzymes).

Biochemistry: The Chemical Composition of Living Matter (pp. 38–55)

1. Inorganic compounds

a. Inorganic compounds making up living matter do not contain carbon (exceptions include CO_2 and CO). They include water, salts, and some acids and bases.

b. Water is the single most abundant compound in the body. It acts as a universal solvent in which electrolytes (salts, acids, and bases) ionize and in which chemical reactions occur, and it is the basis of transport and lubricating fluids. It slowly absorbs and releases heat, thus helping to maintain homeostatic body temperature, and it protects certain body structures (such as the brain) by forming a watery cushion. Water is also a reactant in hydrolysis reactions.

c. Salts in ionic form (electrolytes) are involved in nerve transmission, muscle contraction, blood clotting, transport of oxygen by hemoglobin, metabolism, and many other reactions. Additionally, calcium salts (as bone salts) contribute to bone hardness.

d. Acids are proton donors. When dissolved in water, they release hydrogen ions. Strong acids dissociate completely; weak acids dissociate incompletely.

e. Bases are proton acceptors. The most important inorganic bases are hydroxides. Bicarbonate ions are important bases in the body that act as buffers. When bases and acids interact, neutralization occurs—that is, a salt and water are formed.

f. The relative concentrations of hydrogen and hydroxide ions in various body fluids are measured using a pH scale. Each change of one pH unit represents a tenfold change in hydrogen ion concentration. A pH of 7 is neutral (that is, the concentrations of hydrogen and hydroxide ions are equal). A pH below 7 is acidic; a pH above 7 is alkaline (basic).

g. Normal blood pH ranges from 7.35 to 7.45. Slight deviations outside this range can be fatal.

2. Organic compounds

a. Organic compounds are the carbon-containing compounds that comprise living matter. Carbohydrates, lipids, proteins, and nucleic acids all contain carbon, oxygen, and hydrogen. Proteins and nucleic acids also contain substantial amounts of nitrogen, and nucleic acids also contain phosphorus.

b. Carbohydrates contain carbon, hydrogen, and oxygen in the general relationship of two hydrogen atoms to one oxygen atom and one carbon atom. Their building blocks are monosaccharides. Monosaccharides include glucose, fructose, galactose, deoxyribose, and ribose. Disaccharides include sucrose, maltose, and lactose; and polysaccharides include starch and glycogen. Carbohydrates are ingested as sugars and starches. Carbohydrates, and in particular glucose, are the major energy source for the formation of ATP.

c. Lipids include triglycerides (glycerol plus three fatty acid chains), phospholipids, and steroids (the most important of which is cholesterol). Triglycerides (neutral fats) are found primarily in fatty tissue, where they provide insulation and reserve body fuel. Phospholipids and cholesterol are found in all cell membranes. Cholesterol also forms the basis of certain hormones, bile salts, and vitamin D. Like carbohydrates, the lipids are degraded by hydrolysis and synthesized by dehydration synthesis.

d. Proteins are constructed from building blocks called amino acids; 20 amino acids are found in body proteins.

e. Levels of protein structure include the amino acid sequence (primary); the alpha helix and beta-pleated sheet (secondary); a three-dimensional structure superimposed on secondary structure(s) (tertiary); and a globular structure formed by two or more polypeptide chains (quaternary). Different amino acid sequences result in the construction of different proteins.

f. Fibrous, or structural, proteins are the basic structural materials of the body. Globular proteins are also called functional proteins; examples of these include enzymes, some hormones, and hemoglobin. Functional proteins become denatured and inactivated when their hydrogen bonds are disrupted.

g. Enzymes increase the rates of chemical reactions by binding temporarily and specifically with the reactants and holding them in the proper position to interact. Enzymes do not become part of the product. Many enzymes are produced in an inactive form or are inactivated immediately after use.

h. Nucleic acids include deoxyribonucleic acid (DNA) and ribonucleic acid (RNA). The monomer of nucleic acids is the nucleotide; each nucleotide consists of a nitrogen-containing base, a sugar (ribose or deoxyribose), and a phosphate group. DNA (the "stuff" of the genes) maintains genetic heritage by replicating itself before cell division and specifying protein structure. RNA executes the instructions of the DNA during protein synthesis.

i. ATP (adenosine triphosphate) is the universal energy compound used by all body cells. Some of the energy liberated by the oxidation of glucose is captured in the high-energy phosphate bonds of ATP molecules and stored for later use. Some liberated energy is lost as heat.

Review Questions

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Multiple Choice

More than one choice may apply.

1. Which of the following is (are) true concerning the atomic nucleus?
   a. It contains the mass of the atom.
   b. The negatively charged subatomic particles are here.
   c. Subatomic particles can be ejected.
   d. It contains subatomic particles that determine atomic number.
   e. It contains subatomic particles that interact with other atoms.

2. Pick out the correct match(es) of element and number of valence electrons. Draw a planetary model of each atom to help you choose the best answer.
   a. Oxygen—6
   b. Chlorine—8
   c. Phosphorus—5
   d. Nitrogen—3
   e. Carbon—4

3. Important functions of water include which of the following?
   a. Provides cushioning
   b. Acts as a transport medium
   c. Participates in chemical reactions
   d. Acts as a solvent for sugars, salts, and other solutes
   e. Reduces temperature fluctuations

4. Alkaline substances include which of the following?
   a. Gastric juice
   b. Water
   c. Blood
   d. Lemon juice
   e. Ammonia

5. Glucose is to starch as
   a. a steroid is to a lipid.
   b. a nucleotide is to a nucleic acid.
   c. an amino acid is to a protein.
   d. a polypeptide is to an amino acid.

6. What lipid type is stored in fat deposits beneath the skin?
   a. Triglyceride
   b. Steroid
   c. Vitamin D
   d. Phospholipid
   e. Prostaglandin

7. Absence of which of the following nitrogen-containing bases would prevent RNA synthesis?
   a. Adenine
   b. Cytosine
   c. Guanine
   d. Thymine
   e. Uracil

8. ATP is *not* associated with
   a. a basic nucleotide structure.
   b. high-energy phosphate bonds.
   c. deoxyribose.
   d. inorganic phosphate.
   e. reversible reactions.
9. The element essential for normal thyroid function is
   a. iodine.
   b. iron.
   c. copper.
   d. selenium.
   e. zinc.
10. Factors that increase the speed of chemical reactions include
    a. increasing the temperature.
    b. increasing the particle size.
    c. increasing the concentration of the reactants.
    d. catalysts.

Short Answer Essay

11. Why is a study of basic chemistry essential to understanding human physiology?
12. Matter occupies space and has mass. Explain how energy *must be* described in terms of these two factors. Then define *energy*.
13. Explain the relationship between kinetic and potential energy.
14. Identify the energy form in use in each of the following examples:
    a. Chewing food
    b. Sending a nerve impulse (two types, please—think!)
    c. Bending the fingers to make a fist
    d. Breaking the bonds of ATP molecules to energize your muscle cells to make that fist
15. According to Greek history, a Greek scientist went running through the streets announcing that he had transformed lead into gold. Both lead and gold are elements. On the basis of what you know about the nature of elements, explain why his rejoicing was short-lived.
16. Name and provide the atomic symbols of the four elements that make up the bulk of all living matter. Which of these is found primarily in proteins and nucleic acids?
17. All atoms are neutral. Explain the basis of this fact.
18. Fill in the following table to fully describe an atom's subatomic particles.

Particle	Position in the atom	Charge	Mass
Proton			
Neutron			
Electron			

19. Define *radioactivity*. If an element has three isotopes, which of them (the lightest, the one with an intermediate mass, or the heaviest) is most likely to be a radioisotope, and why?
20. Define *molecule* as it relates to molecular substances and compounds.
21. Explain the basis of ionic bonding.
22. What are hydrogen bonds, and how are they important in the body?
23. The two oxygen atoms forming molecules of oxygen gas that you breathe are joined by a polar covalent bond. Explain why this statement is true or false.
24. Oxygen and argon are both gases. Oxygen combines readily with other elements, but argon does not. What accounts for this difference?
25. Identify each of the following reactions as a synthesis, decomposition, or exchange reaction:

$$2Hg + O_2 \rightarrow 2HgO$$
$$Fe^{2+} + CuSO_4 \rightarrow FeSO_4 + Cu^{2+}$$
$$HCl + NaOH \rightarrow NaCl + H_2O$$
$$HNO_3 \rightarrow H^+ + NO_3^-$$

26. Distinguish inorganic from organic compounds, and list the major categories of each in the body.
27. Salts, acids, and bases are electrolytes. What is an electrolyte?
28. Define *pH*. The pH range of blood is from 7.35 to 7.45. Circle the correct answer to complete the sentence: This is slightly (acidic / basic).
29. A pH of 3.3 is (1 / 10 / 100 / 1000) times more acidic than a pH of 4.3.
30. Define *monosaccharide*, *disaccharide*, and *polysaccharide*. Give at least two examples of each. What is the primary function of carbohydrates in the body?

31. What are the general structures of triglycerides, phospholipids, and steroids? Give one or two important uses of each of these lipid types in the body.

32. The building block of proteins is the amino acid. Draw a diagram of the structure of a generalized amino acid. What is the importance of the R-group?

33. Name the two protein classes based on structure and function in the body, and give two examples of each.

34. Define *enzyme*, and describe enzyme action.

35. Virtually no chemical reaction can occur in the body in the absence of enzymes. How might excessively high body temperature interfere with enzyme activity?

36. What is the structural unit of nucleic acids? Name the two major classes of nucleic acid found in the body, and then compare and contrast them in terms of (a) general three-dimensional structure, and (b) relative functions.

37. What is ATP's central role in the body?

38. Explain why you can "stack" water slightly above the rim of a glass if you pour the water in very carefully.

39. Water is a precious natural resource in California, and it is said that supplies are dwindling. Desalinization (salt removal) of ocean water has been recommended as a solution to the problem. Why shouldn't we drink salt water?

40. Explain the unique chemical makeup and characteristics of phospholipids and what they have to do with the terms *hydrophilic* and *hydrophobic*.

Critical Thinking and Clinical Application Questions

41. Several antibiotics act by binding to certain essential enzymes in the target bacteria. How might these antibiotics influence the chemical reaction controlled by the enzyme? What might be the effect on the bacteria? On the person taking the antibiotic prescription?

42. Mrs. Roberts, who is in a diabetic coma, has just been admitted to Noble Hospital. Her blood pH indicates that she is in severe acidosis (blood pH in the acid range), and the medical staff quickly institute measures to bring her blood pH back within normal limits. Note the normal pH of blood, and discuss why severe acidosis is a problem.

43. Sarah is quite proud of her slender, model-like figure and boasts that she doesn't have an "ounce of excess body fat." Lauren, in contrast, is grossly overweight. She complains of feeling hot most of the time, and on a hot day she is miserable. Sarah generally feels chilled except on very hot days. Explain the relative sensitivity to environmental temperature of these two women on the basis of information you have been given in the Organic Compounds section of this chapter.

44. Pediatricians become concerned about the potential for brain damage when an infant's temperature approaches 105°F. Which class of organic molecules is most likely to be damaged by high temperature? Explain why.

45. The genetic error that causes sickle-cell anemia begins at the DNA level and results in the synthesis of hemoglobin with one amino acid difference compared to normal hemoglobin. Explain how changing one amino acid could affect hemoglobin function.