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Chemistry Comes Alive

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Why study chemistry in an anatomy and physiology course? The answer is simple. Your entire body is made up of chemicals, thousands of them, continuously interacting with one another at an incredible pace. Although it is possible to study anatomy without much reference to chemistry, chemical reactions underlie all physiological processes—movement, digestion, the pumping of your heart, and even your thoughts. This chapter presents the basic chemistry and biochemistry (the chemistry of living material) you need to understand body functions.

PART 1 Basic Chemistry

Definition of Concepts: Matter and Energy

- ✓ Differentiate between matter and energy and between potential energy and kinetic energy.
- ✓ Describe the major energy forms.

Matter

Matter is the “stuff” of the universe. More precisely, **matter** is anything that occupies space and has mass. With some exceptions, it can be seen, smelled, and felt.

For all practical purposes, we can consider mass to be the same as weight. However, this statement is not quite accurate. The *mass* of an object is equal to the actual amount of matter in the object, and it remains constant wherever the object is. In contrast, weight varies with gravity. So while your mass is the same at sea level and on a mountaintop, you weigh just slightly less on that mountaintop. The science of chemistry studies the nature of matter, especially how its building blocks are put together and interact.

States of Matter

Matter exists in *solid*, *liquid*, and *gaseous states*. Examples of each state are found in the human body. Solids, like bones and teeth, have a definite shape and volume. Liquids such as blood plasma have a definite volume, but they conform to the shape of their container. Gases have neither a definite shape nor a definite volume. The air we breathe is a gas.

Energy

Compared with matter, energy is less tangible. It has no mass, does not take up space, and we can measure it only by its effects on matter. **Energy** is defined as the capacity to do work, or to put matter into motion. The greater the work done, the more energy is used doing it. A baseball player who has just hit the ball over the fence uses much more energy than a batter who bunts the ball back to the pitcher.

Kinetic Versus Potential Energy

Energy exists in two forms, or work capacities, and each can be transformed to the other. **Kinetic energy** (ki-net'ik) is energy in action. We see evidence of kinetic energy in the constant movement of the tiniest particles of matter (atoms) as well as in larger objects (a bouncing ball). Kinetic energy does work by moving objects, which in turn can do work by moving or pushing on other objects. For example, a push on a swinging door sets it into motion.

Potential energy is stored energy, that is, inactive energy that has the *potential*, or capability, to do work but is not presently doing so. The batteries in an unused toy have potential energy, as does water confined behind a dam. Your leg muscles have potential energy when you sit still on the couch. When potential energy is released, it becomes kinetic energy and so is capable of doing work. For example, dammed water becomes a rushing torrent when the dam is opened, and that rushing torrent can move a turbine at a hydroelectric plant, or charge a battery.

Actually, energy is a topic of physics, but matter and energy are inseparable. Matter is the substance, and energy is the mover of the substance. All living things are composed of matter and they all require energy to grow and function. The release and use of energy by living systems gives us the elusive quality we call life. Now let's consider the forms of energy used by the body as it does its work.

Forms of Energy

- **Chemical energy** is the form stored in the bonds of chemical substances. When chemical reactions occur that rearrange the atoms of the chemicals in a certain way, the potential energy is unleashed and becomes kinetic energy, or energy in action.

For example, some of the energy in the foods you eat is eventually converted into the kinetic energy of your moving arm. However, food fuels cannot be used to energize body activities directly. Instead, some of the food energy is captured temporarily in the bonds of a chemical called *adenosine triphosphate (ATP)* (ah-den'o-sen tri'fos'fat). Later, ATP's bonds are broken and the stored energy is released as needed to do cellular work. Chemical energy in the form of ATP is the most useful form of energy in living systems because it is used to run almost all functional processes.

- **Electrical energy** results from the movement of charged particles. In your home, electrical energy is found in the flow of electrons along the household wiring. In your body, electrical currents are generated when charged particles called *ions* move along or across cell membranes. The nervous system uses electrical currents, called *nerve impulses*, to transmit messages from one part of the body to another. Electrical currents traveling across the heart stimulate it to contract (beat) and pump blood. (This is why a strong electrical shock, which interferes with such currents, can cause death.)
- **Mechanical energy** is energy *directly* involved in moving matter. When you ride a bicycle, your legs provide the mechanical energy that moves the pedals.
- **Radiant energy**, or **electromagnetic energy** (e-lek'tro-magnet'ik), is energy that travels in waves. These waves, which vary in length, are collectively called the *electromagnetic spectrum*. They include visible light, infrared waves, radio waves, ultraviolet waves, and X rays. Light energy, which stimulates the retinas of our eyes, is important in vision. Ultraviolet waves cause sunburn, but they also stimulate your body to make vitamin D.

Energy Form Conversions

With few exceptions, energy is easily converted from one form to another. For example, the chemical energy (in gasoline) that powers the motor of a speedboat is converted into the mechanical energy of the whirling propeller that makes the boat skim across the water.

Energy conversions are quite inefficient. Some of the initial energy supply is always “lost” to the environment as heat. (It is not really lost because energy cannot be created or destroyed, but that portion given off as heat is at least partly *unusable*.) It is easy to demonstrate this principle. Electrical energy is converted into light energy in a lightbulb. But if you touch a lit bulb, you will soon discover that some of the electrical energy is producing heat instead.

Likewise, all energy conversions in the body liberate heat. This heat helps to maintain our relatively high body temperature, which influences body functioning. For example, when

matter is heated, the kinetic energy of its particles increases and they begin to move more quickly. The higher the temperature, the faster the body's chemical reactions occur. We will learn more about this later.

✓ Check Your Understanding

1. What form of energy is found in the food we eat?
2. What form of energy is used to transmit messages from one part of the body to another?
3. What type of energy is available when we are still? When we are exercising?

For answers, see Appendix H.

Composition of Matter: Atoms and Elements

- ✓ Define chemical element and list the four elements that form the bulk of body matter.
- ✓ Define atom. List the subatomic particles, and describe their relative masses, charges, and positions in the atom.
- ✓ Define atomic number, atomic mass, atomic weight, isotope, and radioisotope.

All matter is composed of **elements**, unique substances that cannot be broken down into simpler substances by ordinary chemical methods. Among the well-known elements are oxygen, carbon, gold, silver, copper, and iron.

At present, 118 elements are recognized. Of these, 92 occur in nature. The rest are made artificially in particle accelerator devices.

Four elements—carbon, oxygen, hydrogen, and nitrogen—make up about 96% of body weight, and 20 others are present in the body, some in trace amounts. **Table 2.1** on p. 26 lists the elements contributing to body mass and gives their importance. In Appendix E, an oddly shaped checkerboard called the **periodic table** provides a more complete listing of the known elements and helps to explain the properties of each element that make it react as it does with other elements.

Each element is composed of more or less identical particles or building blocks, called **atoms**. The smallest atoms are less than 0.1 nanometer (nm) in diameter, and the largest are only about five times as large. [1 nm = 0.0000001 (or 10^{-7}) centimeter (cm), or 40 billionths of an inch!]

Every element's atoms differ from those of all other elements and give the element its unique physical and chemical properties. *Physical properties* are those we can detect with our senses (such as color and texture) or measure (such as boiling point and freezing point). *Chemical properties* pertain to the way atoms interact with other atoms (bonding behavior) and account for the facts that iron rusts, animals can digest their food, and so on.

We designate each element by a one- or two-letter chemical shorthand called an **atomic symbol**, usually the first letter(s) 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

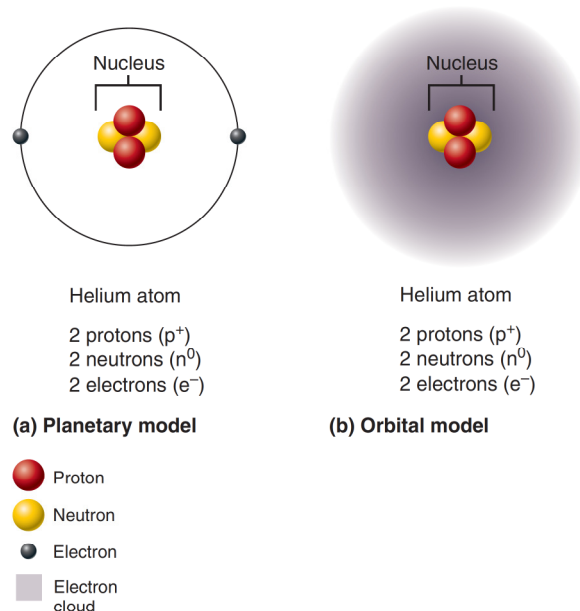


Figure 2.1 Two models of the structure of an atom.

is taken from the Latin name for the element. For example, sodium is indicated by Na, from the Latin word *natrium*.

Atomic Structure

The word *atom* comes from the Greek word meaning “indivisible.” However, we now know that atoms are clusters of even smaller particles called protons, neutrons, and electrons and that even those subatomic particles can be subdivided with high-technology tools. Still, the old idea of atomic indivisibility is useful because an atom loses the unique properties of its element when it is split into its subatomic particles.

An atom's subatomic particles differ in mass, electrical charge, and position in the atom. An atom has a central **nucleus** containing protons and neutrons tightly bound together. The nucleus, in turn, is surrounded by orbiting electrons (**Figure 2.1**). **Protons** (p^+) bear a positive electrical charge, and **neutrons** (n^0) are neutral, so the nucleus is positively charged overall. Protons and neutrons are heavy particles and have approximately the same mass, arbitrarily designated as 1 **atomic mass unit** (1 amu). Since all of the heavy subatomic particles are concentrated in the nucleus, the nucleus is fantastically dense. It accounts for nearly the entire mass (99.9%) of the atom.

The tiny **electrons** (e^-) bear a negative charge equal in strength to the positive charge of the proton. However, an electron has only about 1/2000 the mass of a proton, and the mass of an electron is usually designated as 0 amu.

All atoms are electrically neutral because the number of protons in an atom is precisely balanced by its number of electrons (the + and - charges will then cancel the effect of each other). For example, hydrogen has one proton and one electron, and iron has 26 protons and 26 electrons. For any atom, the number of protons and electrons is always equal.

Table 2.1 Common Elements Composing the Human Body*

ELEMENT	ATOMIC SYMBOL	APPROX. % BODY MASS [†]	FUNCTIONS
Major (96.1%)			
Oxygen	O	65.0	A major component of both organic (carbon-containing) and inorganic (non-carbon-containing) molecules. As a gas, it is needed for the production of cellular energy (ATP).
Carbon	C	18.5	A primary component of all organic molecules, which include carbohydrates, lipids (fats), proteins, and nucleic acids.
Hydrogen	H	9.5	A component of all organic molecules. As an ion (proton), 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. Its ionic (Ca^{2+}) form is required for muscle contraction, conduction of nerve impulses, and blood clotting.
Phosphorus	P	1.0	Part of calcium phosphate salts in bones and teeth. Also present in nucleic acids, and part of ATP.
Potassium	K	0.4	Its ion (K^+) is the major positive ion (cation) in cells. Necessary for conduction of nerve impulses and muscle contraction.
Sulfur	S	0.3	Component of proteins, particularly muscle proteins.
Sodium	Na	0.2	As an ion (Na^+), sodium is the major positive ion found in extracellular fluids (fluids outside of cells). Important for water balance, conduction of nerve impulses, and muscle contraction.
Chlorine	Cl	0.2	Its ion (chloride, Cl^-) is the most abundant negative ion (anion) in extracellular fluids.
Magnesium	Mg	0.1	Present in bone. Also an important cofactor in a number of metabolic reactions.
Iodine	I	0.1	Needed to make functional thyroid hormones.
Iron	Fe	0.1	Component of hemoglobin (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)			
These elements are referred to as <i>trace elements</i> because they are required in very minute amounts; many are found as part of enzymes or are required for enzyme activation.			

*A listing of the elements by ascending order of atomic number appears in the periodic table, Appendix E.

[†]Percentage of “wet” body mass; includes water.

The **planetary model** of the atom, illustrated in Figure 2.1a, is a simplified (and now outdated) model of atomic structure. As you can see, it depicts electrons moving around the nucleus in fixed, generally circular orbits. But we can never determine the exact location of electrons at a particular time because they jump around following unknown trajectories. So, instead of speaking of specific orbits, chemists talk about **orbitals**—regions around the nucleus in which a given electron or electron pair is likely to be found most of the time. This more modern model of atomic structure, called the **orbital model**, is more useful for predicting the chemical behavior of atoms. As illustrated in Figure 2.1b, the orbital model depicts *probable* regions of greatest electron density by denser shading (this haze is called

the *electron cloud*). However, the planetary model is simpler to depict, so we will use that model in most illustrations of atomic structure in this text.

Hydrogen, with just one proton and one electron, is the simplest atom. You can visualize the spatial relationships in the hydrogen atom by imagining it as a sphere enlarged until its diameter equals the length of a football field. In that case, the nucleus could be represented by a lead ball the size of a gumdrop in the exact center of the sphere. Its lone electron could be pictured as a fly buzzing about unpredictably within the sphere. Though not completely accurate, this mental image demonstrates that most of the volume of an atom is empty space, and nearly all of its mass is concentrated in the central nucleus.

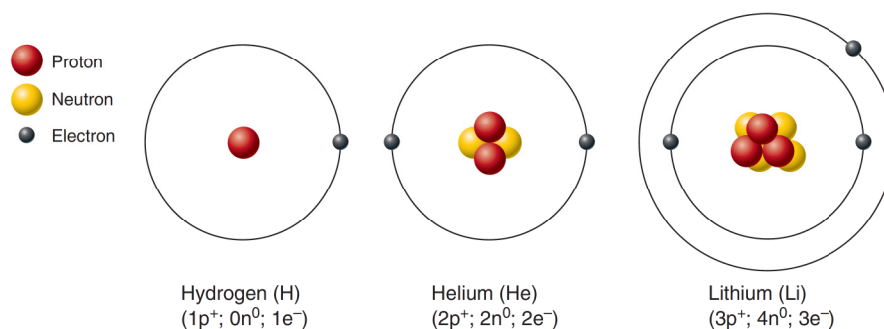


Figure 2.2 Atomic structure of the three smallest atoms.

Identifying Elements

All protons are alike, regardless of the atom considered. The same is true of all neutrons and all electrons. So what determines the unique properties 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 in size 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 continued this step-by-step progression, we would get a graded series of atoms containing from 1 to 112 protons, an equal number of electrons, and a slightly larger number of neutrons at each step.

All we really need to know to identify a particular element, however, are its atomic number, mass number, and atomic weight. Taken together, these provide a fairly complete picture of each element.

Atomic Number

The **atomic number** of any atom is equal to the number of protons in its nucleus and is written as a subscript to the left of its atomic symbol. Hydrogen, with one proton, has an atomic number of 1 (${}^1\text{H}$). Helium, with two protons, has an atomic number of 2 (${}^2\text{He}$), and so on. The number of protons is always equal to the number of electrons in an atom, so the atomic number *indirectly* tells us the number of electrons in the atom as well. As we will see shortly, this information is important indeed, because electrons determine the chemical behavior of atoms.

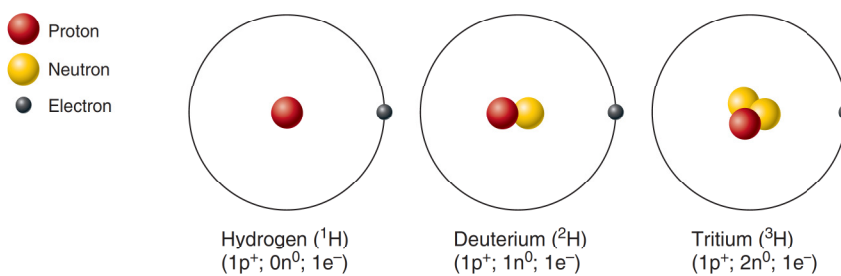


Figure 2.3 Isotopes of hydrogen.

Mass Number and Isotopes

The **mass number** of an atom is the sum of the masses of its protons and neutrons. The mass of the electrons is so small that it is ignored. Recall that protons and neutrons have a mass of 1 amu. Hydrogen has only one proton in its nucleus, so its atomic and mass numbers are the same: 1. Helium, with two protons and two neutrons, has a mass number of 4.

The mass number is usually indicated by a superscript to the left of the atomic symbol. For example, helium is ${}^4\text{He}$. This simple notation allows us to deduce the total number and kinds of subatomic particles in any atom because it indicates the number of protons (the atomic number), the number of electrons (equal to the atomic number), and the number of neutrons (mass number minus atomic number). In our example, we can do the subtraction to find that ${}^4\text{He}$ has two neutrons.

From what we have said so far, it may appear as if each element has one, and only one, type of atom representing it. This is not the case. Nearly all known elements have two or more structural variations called **isotopes** (iso-tōps), which have the same number of protons (and electrons), but differ in the number of neutrons they contain. Earlier, when we said that hydrogen has a mass number of 1, we were speaking of ${}^1\text{H}$, its most abundant isotope. Some hydrogen atoms have a mass of 2 or 3 amu (atomic mass units), which means that they have one proton and, respectively, one or two neutrons (**Figure 2.3**).

Carbon has several isotopes. The most abundant of these are ^{12}C , ^{13}C , and ^{14}C . Each of the carbon isotopes has six protons (otherwise it would not be carbon), but ^{12}C has six neutrons, ^{13}C has seven, and ^{14}C has eight. Isotopes can also be written with the mass number following the symbol: C-14, for example.

Atomic Weight

You might think that atomic weight should be the same as atomic mass, and this would be so if atomic weight referred to the weight of a single atom. However, **atomic weight** is an average of the relative weights (mass numbers) of *all* the isotopes of an element, taking into account their relative abundance in nature. As a rule, the atomic weight of an element is approximately equal to the mass number of its most abundant isotope. For example, the atomic weight of hydrogen is 1.008, which reveals that its lightest isotope (^1H) is present in much greater amounts in our world than its ^2H or ^3H forms.

Radioisotopes

The heavier isotopes of many elements are unstable, and their atoms decompose spontaneously into more stable forms. This process of atomic decay is called *radioactivity*, and isotopes that exhibit this behavior are called **radioisotopes** (ra"de-o-i"so-tops). The disintegration of a radioactive nucleus may be compared to a tiny explosion. It occurs when subatomic *alpha* (α) *particles* (packets of $2p + 2n$), *beta* (β) *particles* (electron-like negative particles), or *gamma* (γ) *rays* (electromagnetic energy) are ejected from the atomic nucleus.

Why does this happen? The answer is complex, but for our purposes, the important point to know is that the dense nuclear particles are composed of even smaller particles called *quarks* that associate in one way to form protons and in another way to form neutrons. Apparently, the "glue" that holds these nuclear particles together is weaker in the heavier isotopes. When radioisotopes disintegrate, the element may transform to a different element.

Because we can detect radioactivity with scanners, and radioactive isotopes share the same chemistry as their more stable isotopes, radioisotopes are valuable tools for biological research and medicine. Most radioisotopes used in the clinical setting are used for diagnosis, that is, to localize and illuminate damaged or cancerous tissues. For example, iodine-131 is used to determine the size and activity of the thyroid gland and to detect thyroid cancer. The sophisticated PET scans described in *A Closer Look* in Chapter 1 use radioisotopes to probe the workings of molecules deep within our bodies. All radioisotopes, regardless of the purpose for which they are used, damage living tissue, and all radioisotopes gradually lose their radioactive behavior. The time required for a radioisotope to lose one-half of its activity is called its *half-life*. The half-lives of radioisotopes vary dramatically from hours to thousands of years.

Alpha emission is easily blocked outside the body but if absorbed causes considerable damage. For this reason, inhaled alpha particles from decaying radon are second only to smoking

as a cause of lung cancer. (Radon results naturally from decay of uranium in the ground.) Gamma emission has the greatest penetrating power. Radium-226, cobalt-60, and certain other radioisotopes that decay by gamma emission are used to destroy localized cancers.

Contrary to what some believe, ionizing radiation does not damage organic molecules directly. Instead, it knocks electrons out of other atoms and sends them flying, like bowling balls smashing through pins all along their path. It is the electron energy and the unstable molecules left behind that do the damage.

✓ Check Your Understanding

4. What two elements besides H and N make up the bulk of living matter?
5. An element has a mass of 207 and has 125 neutrons in its nucleus. How many protons and electrons does it have and where are they located?
6. How do the terms atomic mass and atomic weight differ?

For answers, see Appendix H.

How Matter Is Combined: Molecules and Mixtures

- ✓ Define molecule, and distinguish between a compound and a mixture.
- ✓ Compare solutions, colloids, and suspensions.

Molecules and Compounds

Most atoms do not exist in the free state, but instead are chemically combined with other atoms. Such a combination of two or more atoms held together by chemical bonds is called a **molecule**.

If two or more atoms of the *same* element combine, the resulting substance is called a *molecule of that element*. When two hydrogen atoms bond, the product is a molecule of hydrogen gas and is written as H_2 . Similarly, when two oxygen atoms combine, a molecule of oxygen gas (O_2) is formed. Sulfur atoms commonly combine to form sulfur molecules containing eight sulfur atoms (S_8).

When two or more *different* kinds of atoms bind, they form molecules of a **compound**. Two hydrogen atoms combine with one oxygen atom to form the compound water (H_2O). Four hydrogen atoms combine with one carbon atom to form the compound methane (CH_4). Notice again that molecules of methane and water are compounds, but molecules of hydrogen gas are not, because compounds always contain atoms of at least two different elements.

Compounds are chemically pure, and all of their molecules are identical. So, just as an atom is the smallest particle of an element that still has the properties of the element, a molecule is the smallest particle of a compound that still has the specific

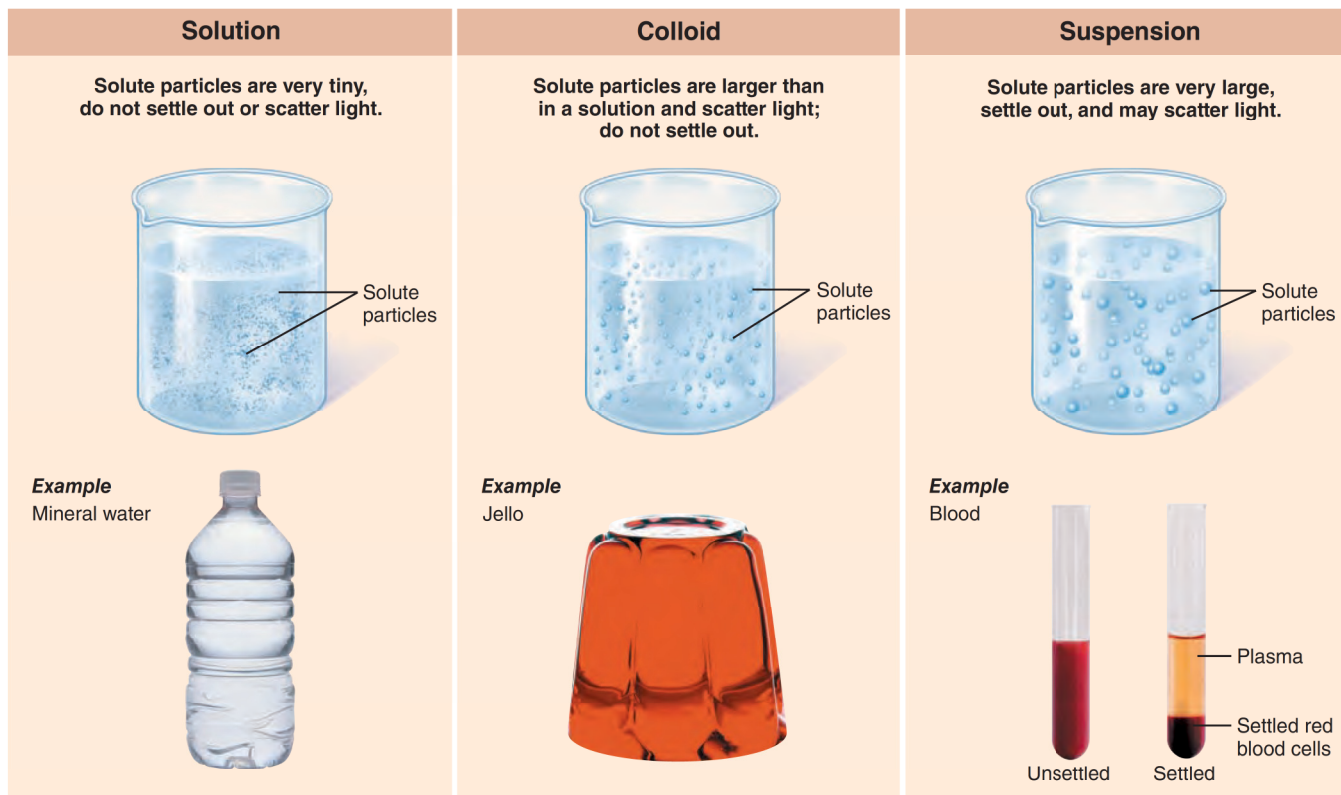


Figure 2.4 The three basic types of mixtures.

characteristics of the compound. This concept is important because the properties of compounds are usually very different from those of the atoms they contain. Water, for example, is very different from the elements hydrogen and oxygen. Indeed, it is next to impossible to tell what atoms are in a compound without analyzing it chemically.

Mixtures

Mixtures are substances composed of two or more components *physically intermixed*. Most matter in nature exists in the form of mixtures, but there are only three basic types: *solutions*, *colloids*, and *suspensions* (Figure 2.4).

Solutions

Solutions are homogeneous mixtures of components that may be gases, liquids, or solids. *Homogeneous* means that the mixture has exactly the same composition or makeup throughout—a sample taken from any part of the mixture has the same composition (in terms of the atoms or molecules it contains) as a sample taken from any other part of the mixture. Examples include the air we breathe (a mixture of gases) and seawater (a mixture of salts, which are solids, and water). The substance present in the greatest amount is called the **solvent** (or dissolving medium). Solvents are usually liquids. Substances present in smaller amounts are called **solutes**.

Water is the body's chief solvent. Most solutions in the body are *true solutions* containing gases, liquids, or solids dissolved in water. True solutions are usually transparent. Examples are saline solution [table salt (NaCl) and water], a mixture of glucose and water, and mineral water. The solutes of true solutions are minute, usually in the form of individual atoms and molecules. Consequently, they are not visible to the naked eye, do not settle out, and do not scatter light. In other words, if a beam of light is passed through a true solution, you will not see the path of light.

Concentration of Solutions We describe true solutions in terms of their *concentration*, which may be indicated in various ways. Solutions used in a college laboratory or a hospital are often described in terms of the **percent** (parts per 100 parts) of the solute in the total solution. This designation always refers to the solute percentage, and unless otherwise noted, water is assumed to be the solvent.

Milligrams per deciliter (mg/dl) is another common concentration measurement. (A deciliter is 100 milliliters or 0.1 liter.)

Still another way to express the concentration of a solution is in terms of its **molarity** (mo-lar'y-te), or moles per liter, indicated by *M*. This method is more complicated but much more useful. To understand molarity, you must understand what a mole is. A **mole** of any element or compound is equal to its atomic weight or **molecular weight** (sum of the atomic weights) weighed out in grams. This concept is easier than it seems, as illustrated by the following example.

Glucose is $C_6H_{12}O_6$, which indicates that it has 6 carbon atoms, 12 hydrogen atoms, and 6 oxygen atoms. To compute the molecular weight of glucose, you would look up the atomic weight of each of its atoms in the periodic table (see Appendix E) and compute its molecular weight as follows:

Atom	Number of Atoms		Atomic Weight	=	Total Atomic Weight
C	6	×	12.011	=	72.066
H	12	×	1.008	=	12.096
O	6	×	15.999	=	95.994
					180.156

Then, to make a *one-molar* solution of glucose, you would weigh out 180.156 grams (g), called a *gram molecular weight*, of glucose and add enough water to make 1 liter (L) of solution. In short, a one-molar solution (abbreviated 1.0 M) of a chemical substance is one gram molecular weight of the substance (or one gram atomic weight in the case of elemental substances) in 1 L (1000 milliliters) of solution.

The beauty of using the mole as the basis of preparing solutions is its precision. One mole of any substance always contains exactly the same number of solute particles, that is, 6.02×10^{23} . This number is called **Avogadro's number** (av"o-gad'rōz). So whether you weigh out 1 mole of glucose (180 g) or 1 mole of water (18 g) or 1 mole of methane (16 g), in each case you will have 6.02×10^{23} molecules of that substance.* This allows almost mind-boggling precision to be achieved.

Because solute concentrations in body fluids tend to be quite low, those values are usually reported in terms of millimoles (mM; 1/1000 mole).

Colloids

Colloids (kol'oidz), also called *emulsions*, are *heterogeneous* mixtures, which means that their composition is dissimilar in different areas of the mixture. Colloids often appear translucent or milky and although the solute particles are larger than those in true solutions, they still do not settle out. However, they do scatter light, so the path of a light beam shining through a colloidal mixture is visible.

Colloids have many unique properties, including the ability of some to undergo **sol-gel transformations**, that is, to change reversibly from a fluid (sol) state to a more solid (gel) state. Jell-O, or any gelatin product (Figure 2.4), is a familiar example of a nonliving colloid that changes from a sol to a gel when refrigerated (and that gel will liquefy again if placed in the sun). Cytosol, the semifluid material in living cells, is also a colloid,

*The important exception to this rule concerns molecules that ionize and break up into charged particles (ions) in water, such as salts, acids, and bases (see p. 39). For example, simple table salt (sodium chloride) breaks up into two types of charged particles. Therefore, in a 1.0 M solution of sodium chloride, 2 moles of solute particles are actually in solution.

largely because of its dispersed proteins. Its sol-gel transformations underlie many important cell activities, such as cell division and changes in cell shape.

Suspensions

Suspensions are *heterogeneous* mixtures with large, often visible solutes that tend to settle out. An example of a suspension is a mixture of sand and water. So is blood, in which the living blood cells are suspended in the fluid portion of blood (blood plasma). If left to stand, the suspended cells will settle out unless some means—mixing, shaking, or circulation in the body—keeps them in suspension.

As you can see, all three types of mixtures are found in both living and nonliving systems. In fact, living material is the most complex mixture of all, since it contains all three kinds of mixtures interacting with one another.

Distinguishing Mixtures from Compounds

Now let's zero in on how to distinguish mixtures and compounds from one another. Mixtures differ from compounds in several important ways:

- The chief difference between mixtures and compounds is that no chemical bonding occurs between the components of a mixture. The properties of atoms and molecules are not changed when they become part of a mixture. Remember they are only physically intermixed.
- Depending on the mixture, its components can be separated by physical means—straining, filtering, evaporation, and so on. Compounds, by contrast, can be separated into their constituent atoms only by chemical means (breaking bonds).
- Some mixtures are homogeneous, whereas others are heterogeneous. A bar of 100% pure (elemental) iron is homogeneous, as are all compounds. As already mentioned, heterogeneous substances vary in their makeup from place to place. For example, iron ore is a heterogeneous mixture that contains iron and many other elements.

✓ Check Your Understanding

7. What is the meaning of the term "molecule"?
8. Why is sodium chloride (NaCl) considered a compound, but oxygen gas is not?
9. Blood contains a liquid component and living cells. Would it be classified as a compound or a mixture? Why?

For answers, see Appendix H.

Chemical Bonds

- ✓ Explain the role of electrons in chemical bonding and in relation to the octet rule.
- ✓ Differentiate among ionic, covalent, and hydrogen bonds.
- ✓ Compare and contrast polar and nonpolar compounds.

As noted earlier, when atoms combine with other atoms, they are held together by **chemical bonds**. A chemical bond is not a physical structure like a pair of handcuffs linking two people together. Instead, it is an energy relationship between the electrons of the reacting atoms, and it is made or broken in less than a trillionth of a second.

The Role of Electrons in Chemical Bonding

Electrons forming the electron cloud around the nucleus of an atom occupy regions of space called **electron shells** that consecutively surround the atomic nucleus. The atoms known so far can have electrons in seven shells (numbered 1 to 7 from the nucleus outward), but the actual number of electron shells occupied in a given atom depends on the number of electrons that atom has. Each electron shell contains one or more orbitals. (Recall from our earlier discussion that *orbitals* are regions around the nucleus in which a given electron is likely to be found most of the time.)

It is important to understand that each electron shell represents a different **energy level**, because this prompts you to think of electrons as particles with a certain amount of potential energy. In general, the terms *electron shell* and *energy level* are used interchangeably.

How much potential energy does an electron have? The answer depends on the energy level that it occupies. The attraction between the positively charged nucleus and negatively charged electrons is greatest when electrons are closest to the nucleus and falls off with increasing distance. This statement explains why electrons farthest from the nucleus (1) have the greatest potential energy (it takes more energy for them to overcome the nuclear attraction and reach the more distant energy levels) and (2) are most likely to interact chemically with other atoms. (They are the least tightly held by their own atomic nucleus and the most easily influenced by other atoms and molecules.)

Each electron shell can hold a specific number of electrons. Shell 1, the shell immediately surrounding the nucleus, accommodates only 2 electrons. Shell 2 holds a maximum of 8, and shell 3 has room for 18. Subsequent shells hold larger and larger numbers of electrons, and the shells tend to be filled with electrons consecutively. For example, shell 1 fills completely before any electrons appear in shell 2.

Which electrons are involved in chemical bonding? When we consider bonding behavior, the only electrons that are important are those in the atom's outermost energy level. Inner electrons usually do not take part in bonding because they are more tightly held by the atomic nucleus.

When the outermost energy level of an atom is filled to capacity or contains eight electrons, the atom is stable. Such atoms are *chemically inert*, that is, unreactive. A group of elements called the *noble gases*, which include helium and neon, typify this condition (Figure 2.5a). On the other hand, atoms in which the outermost energy level contains fewer than eight electrons tend to gain, lose, or share electrons with other atoms to achieve stability (Figure 2.5b).

What about atoms that have more than 20 electrons, in which the energy levels beyond shell 2 can contain *more* than

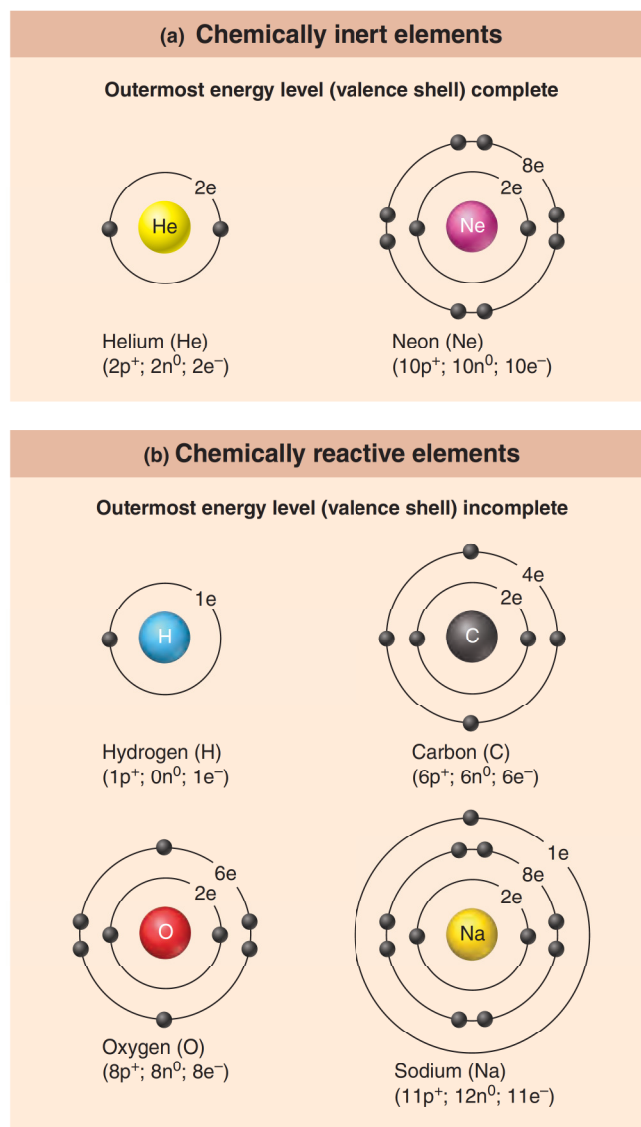


Figure 2.5 Chemically inert and reactive elements. (Note: For simplicity, each atomic nucleus is shown as a sphere with the atom's symbol; individual protons and neutrons are not shown.)

eight electrons? The number of electrons that can participate in bonding is still limited to a total of eight. The term **valence shell** (va'lens) specifically indicates an atom's outermost energy level *or that portion of it* containing the electrons that are chemically reactive. Hence, the key to chemical reactivity is the **octet rule** (ok-tet'), or **rule of eights**. Except for shell 1, which is full when it has two electrons, atoms tend to interact in such a way that they have eight electrons in their valence shell.

Types of Chemical Bonds

Three major types of chemical bonds—*ionic*, *covalent*, and *hydrogen bonds*—result from attractive forces between atoms.

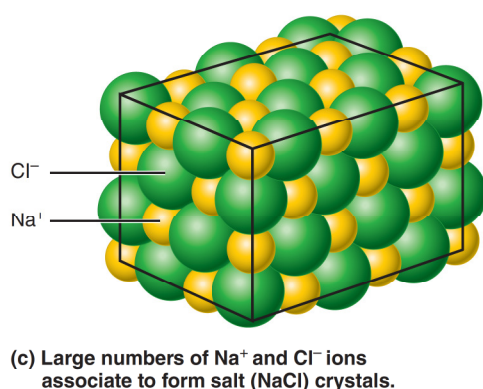
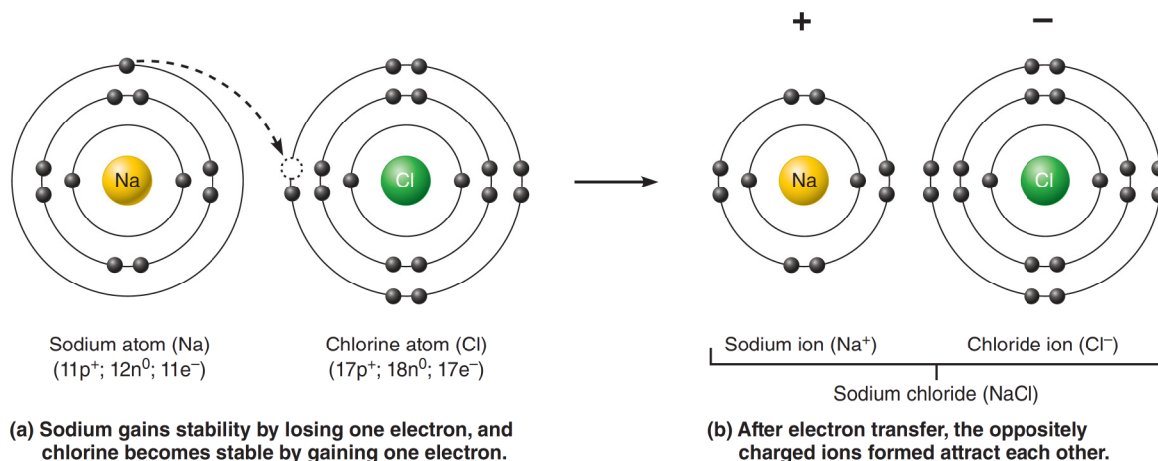


Figure 2.6 Formation of an ionic bond.

Ionic Bonds

Recall that atoms are electrically neutral. However, electrons can be transferred from one atom to another, and when this happens, the precise balance of + and - charges is lost so that charged particles called **ions** are formed. An **ionic bond** (i-on'ik) is a chemical bond between atoms formed by the transfer of one or more electrons from one atom to the other. The atom that gains one or more electrons is the *electron acceptor*. It acquires a net negative charge and is called an **anion** (an'i-on). The atom that loses electrons is the *electron donor*. It acquires a net positive charge and is called a **cation** (kat'i-on). (To remember this term, think of the "t" in "cation" as a + sign.) Both anions and cations are formed whenever electron transfer between atoms occurs. Because opposite charges attract, these ions tend to stay close together, resulting in an ionic bond.

One example of ionic bonding is the formation of table salt, or sodium chloride (NaCl), by interaction of sodium and chlorine atoms (**Figure 2.6**). Sodium, with an atomic number of 11, has only one electron in its valence shell. It would be very difficult to attempt to fill this shell by adding seven more. However, if this single electron is lost, shell 2 with eight electrons becomes the valence shell (outermost energy level containing electrons) and is full. Thus, by losing the lone electron in its third energy

level, sodium achieves stability and becomes a cation (Na^+). On the other hand, chlorine, atomic number 17, needs only one electron to fill its valence shell. By accepting an electron, chlorine achieves stability and becomes an anion.

When sodium and chlorine atoms interact, this is exactly what happens. Sodium donates an electron to chlorine (**Figure 2.6a**), and the oppositely charged ions created in this exchange attract each other, forming sodium chloride (**Figure 2.6b**). Ionic bonds are commonly formed between atoms with one or two valence shell electrons (the metallic elements, such as sodium, calcium, and potassium) and atoms with seven valence shell electrons (such as chlorine, fluorine, and iodine).

Most ionic compounds fall in the chemical category called *salts*. In the dry state, salts such as sodium chloride do not exist as individual molecules. Instead, they form **crystals**, large arrays of cations and anions held together by ionic bonds (**Figure 2.6c**).

Sodium chloride is an excellent example of the difference in properties between a compound and its constituent atoms. Sodium is a silvery white metal, and chlorine in its molecular state is a poisonous green gas used to make bleach. However, sodium chloride is a white crystalline solid that we sprinkle on our food.

Covalent Bonds

Electrons do not have to be completely transferred for atoms to achieve stability. Instead, they may be *shared* so that each atom is able to fill its outer electron shell at least part of the time. Electron sharing produces molecules in which the shared electrons occupy a single orbital common to both atoms, which constitutes a **covalent bond** (ko-va'lent).

Hydrogen with its single electron can fill its only shell (shell 1) by sharing a pair of electrons with another atom. When it shares with another hydrogen atom, a molecule of hydrogen gas is formed. The shared electron pair orbits around the molecule as a whole, satisfying the stability needs of each atom.

Hydrogen can also share an electron pair with different kinds of atoms to form a compound (**Figure 2.7a**). Carbon has four electrons in its outermost shell, but needs eight to achieve stability. Hydrogen has one electron, but needs two. When a methane

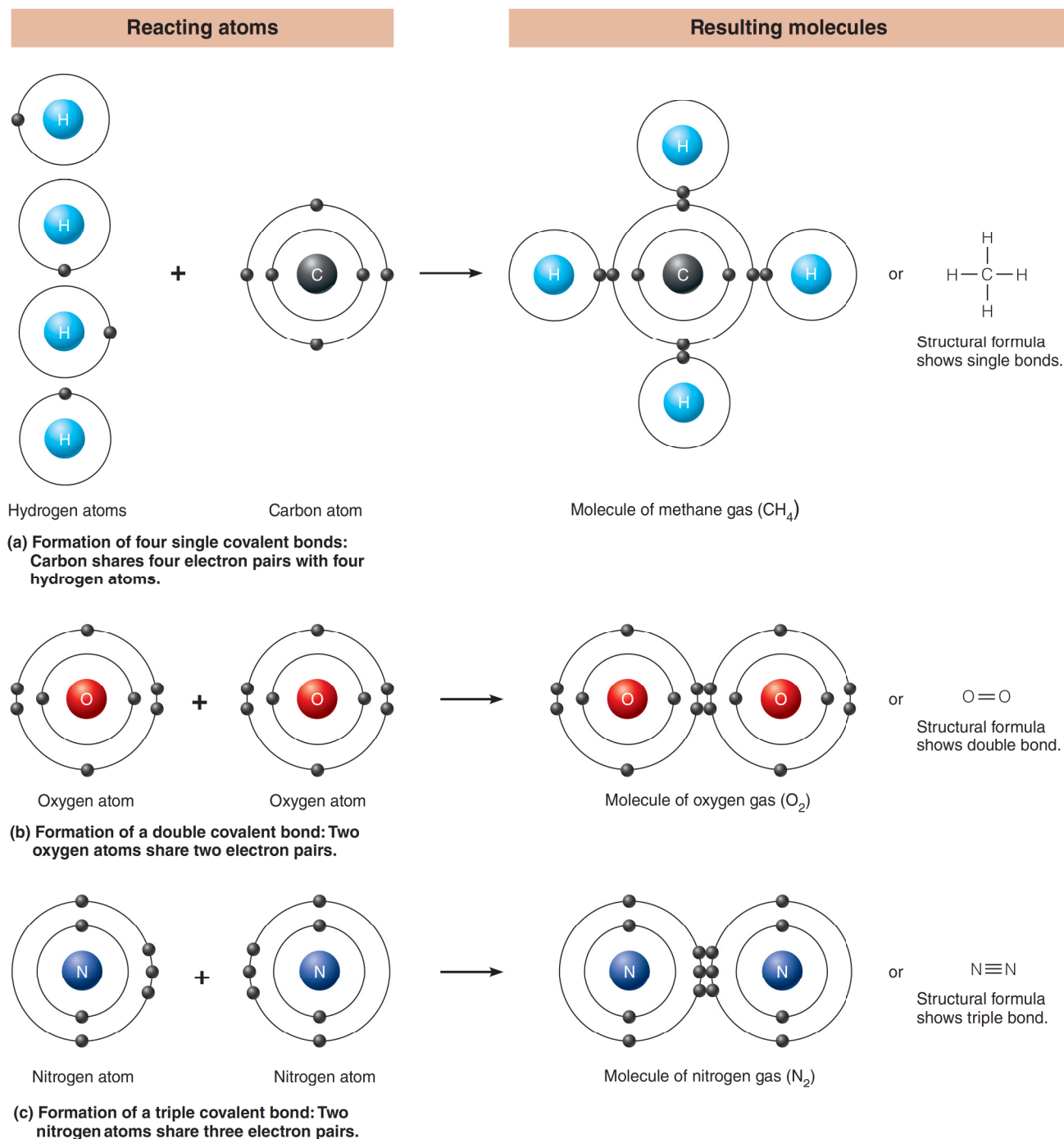


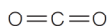
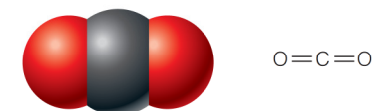
Figure 2.7 Formation of covalent bonds.

molecule (CH₄) is formed, carbon shares four pairs of electrons with four hydrogen atoms (one pair with each hydrogen). Again, the shared electrons orbit and “belong to” the whole molecule, ensuring the stability of each atom.

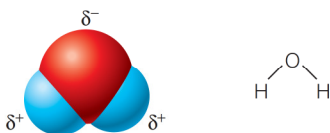
When two atoms share one pair of electrons, a single covalent bond is formed (indicated by a single line connecting the atoms, such as H—H). In some cases, atoms share two or three

electron pairs, resulting in *double* or *triple covalent bonds* (Figure 2.7b and c). (These bonds are indicated by double or triple connecting lines such as O=O or N≡N.)

Polar and Nonpolar Molecules In the covalent bonds we have discussed, the shared electrons are shared equally between the atoms of the molecule for the most part. The molecules formed



(a) Carbon dioxide (CO_2) molecules are linear and symmetrical. They are nonpolar.



(b) V-shaped water (H_2O) molecules have two poles of charge—a slightly more negative oxygen end (δ^-) and a slightly more positive hydrogen end (δ^+).

Figure 2.8 Carbon dioxide and water molecules have different shapes, as illustrated by molecular models.

are electrically balanced and are called **nonpolar molecules** (because they do not have separate + and – poles of charge).

Such electrical balance is not always the case. When covalent bonds are formed, the resulting molecule always has a specific three-dimensional shape, with the bonds formed at definite angles. A molecule's shape helps determine what other molecules or atoms it can interact with. It may also result in unequal electron pair sharing, creating a **polar molecule**, especially in nonsymmetrical molecules containing atoms with different electron-attracting abilities.

In general, *small* atoms with six or seven valence shell electrons, such as oxygen, nitrogen, and chlorine, are electron-hungry and attract electrons very strongly, a capability called **electronegativity**. On the other hand, most atoms with only one or two valence shell electrons tend to be **electropositive**. In other words, their electron-attracting ability is so low that they usually lose *their* valence shell electrons to other atoms. Potassium and sodium, each with one valence shell electron, are good examples of electropositive atoms.

Carbon dioxide and water illustrate how molecular shape and the relative electron-attracting abilities of atoms determine whether a covalently bonded molecule is nonpolar or polar. In carbon dioxide (CO_2), carbon shares four electron pairs with two oxygen atoms (two pairs are shared with each oxygen). Oxygen is very electronegative and so attracts the shared electrons much more strongly than does carbon. However, because the carbon dioxide molecule is linear and symmetrical (**Figure 2.8a**), the electron-pulling ability of one oxygen atom offsets that of the other, like a standoff between equally strong teams in a game of tug-of-war. As a result, the shared electrons orbit the entire molecule and carbon dioxide is a nonpolar compound.

In contrast, a water molecule (H_2O) is bent, or V shaped (**Figure 2.8b**). The two electropositive hydrogen atoms are located at the same end of the molecule, and the very electronegative oxygen is at the opposite end. This arrangement allows oxygen to pull the shared electrons toward itself and away from the two hydrogen atoms. In this case, the electron pairs are *not* shared

←-----→		
Ionic bond	Polar covalent bond	Nonpolar covalent bond
Complete transfer of electrons	Unequal sharing of electrons	Equal sharing of electrons
Separate ions (charged particles) form	Slight negative charge (δ^-) at one end of molecule, slight positive charge (δ^+) at other end	Charge balanced among atoms
$\text{Na}^+ \text{Cl}^-$ Sodium chloride	 Water	$\text{O}=\text{C}=\text{O}$ Carbon dioxide

Figure 2.9 Ionic, polar covalent, and nonpolar covalent bonds compared along a continuum.

equally, but spend more time in the vicinity of oxygen. Because electrons are negatively charged, the oxygen end of the molecule is slightly more negative (the charge is indicated with a delta and minus as δ^-) and the hydrogen end slightly more positive (indicated by δ^+). Because water has two poles of charge, it is a *polar molecule*, or **dipole** (di'pōl).

Polar molecules orient themselves toward other dipoles or toward charged particles (such as ions and some proteins), and they play essential roles in chemical reactions in body cells. The polarity of water is particularly significant, as you will see later in this chapter.

Different molecules exhibit different degrees of polarity, and we can see a gradual change from ionic to nonpolar covalent bonding as summarized in **Figure 2.9**. Ionic bonds (complete electron transfer) and nonpolar covalent bonds (equal electron sharing) are the extremes of a continuum, with various degrees of unequal electron sharing in between.

Hydrogen Bonds

Unlike the stronger ionic and covalent bonds, hydrogen bonds are more like attractions than true bonds. Hydrogen bonds form when a hydrogen atom, already covalently linked to one electronegative atom (usually nitrogen or oxygen), is attracted by another electron-hungry atom, so that a “bridge” forms between them.

Hydrogen bonding is common between dipoles such as water molecules, because the slightly negative oxygen atoms of one molecule attract the slightly positive hydrogen atoms of other molecules (**Figure 2.10a**). Hydrogen bonding is responsible for the tendency of water molecules to cling together and form films, referred to as *surface tension*. This tendency helps explain why water

beads up into spheres when it sits on a hard surface and why water striders can walk on a pond's surface (Figure 2.10b).

Although hydrogen bonds are too weak to bind atoms together to form molecules, they are important *intramolecular bonds* (literally, bonds within molecules), which hold different parts of a single large molecule in a specific three-dimensional shape. Some large biological molecules, such as proteins and DNA, have numerous hydrogen bonds that help maintain and stabilize their structures.

✓ Check Your Understanding

10. What kinds of bonds form between water molecules?
11. Oxygen (${}_{8}\text{O}$) and argon (${}_{18}\text{Ar}$) are both gases. Oxygen combines readily with other elements, but argon does not. What accounts for this difference?
12. Assume imaginary compound XY has a polar covalent bond. How does its charge distribution differ from that of XX molecules?

For answers, see Appendix H.

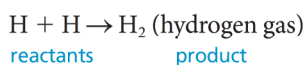
Chemical Reactions

- ✓ Define the three major types of chemical reactions: synthesis, decomposition, and exchange. Comment on the nature of oxidation-reduction reactions and their importance.
- ✓ Explain why chemical reactions in the body are often irreversible.
- ✓ Describe factors that affect chemical reaction rates.

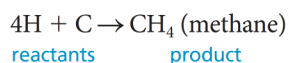
As we noted earlier, all particles of matter are in constant motion because of their kinetic energy. Movement of atoms or molecules in a solid is usually limited to vibration because the particles are united by fairly rigid bonds. But in liquids or gases, particles dart about randomly, sometimes colliding with one another and interacting to undergo chemical reactions. A **chemical reaction** occurs whenever chemical bonds are formed, rearranged, or broken.

Chemical Equations

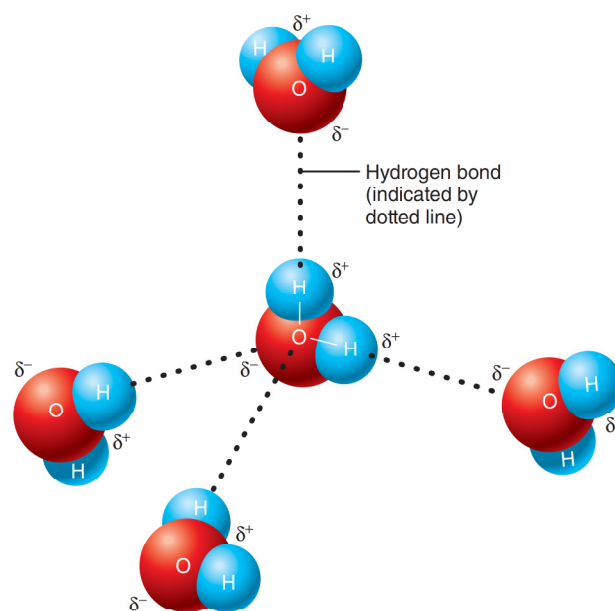
We can write chemical reactions in symbolic form as chemical equations. For example, we indicate the joining of two hydrogen atoms to form hydrogen gas as



and the combining of four hydrogen atoms and one carbon atom to form methane as



Notice that in equations, a number written as a *subscript* indicates that the atoms are joined by chemical bonds. But a number written as a *prefix* denotes the number of *unjoined* atoms or molecules. For example, CH_4 reveals that four hydrogen



(a) The slightly positive ends (δ^+) of the water molecules become aligned with the slightly negative ends (δ^-) of other water molecules.



(b) A water strider can walk on a pond because of the high surface tension of water, a result of the combined strength of its hydrogen bonds.

Figure 2.10 Hydrogen bonding between polar water molecules.

atoms are bonded together with carbon to form the methane molecule, but 4H signifies four unjoined hydrogen atoms.

A chemical equation is like a sentence describing what happens in a reaction. It contains the following information: the number and kinds of reacting substances, or **reactants**; the chemical composition of the **product(s)**; and in balanced equations, the relative proportion of each reactant and product.

In the equations above, the reactants are atoms, as indicated by their atomic symbols (H, C). The product in each case is a

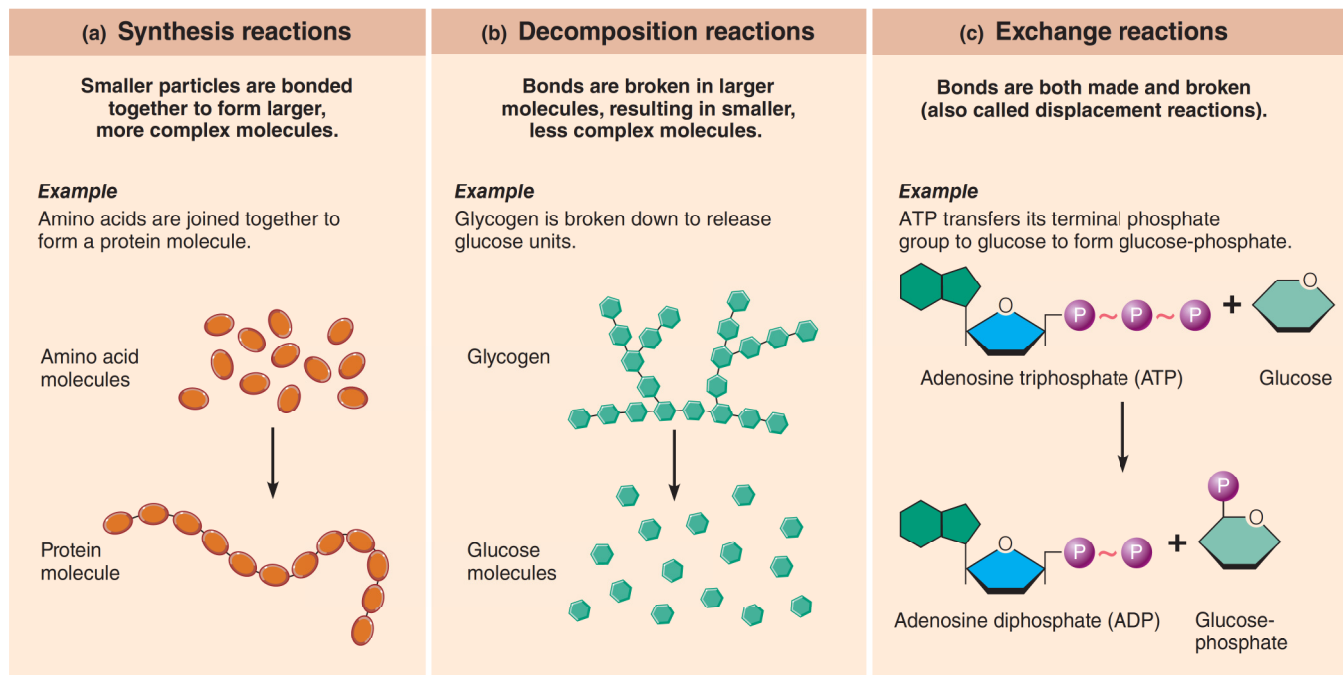


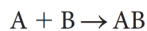
Figure 2.11 Patterns of chemical reactions.

molecule, as represented by its **molecular formula** (H_2 , CH_4). The equation for the formation of methane may be read in terms of molecules or moles—as *either* “four hydrogen atoms plus one carbon atom yield one molecule of methane” or “four moles of hydrogen atoms plus one mole of carbon yield one mole of methane.” Using moles is more practical because it is impossible to measure out one atom or one molecule of anything!

Patterns of Chemical Reactions

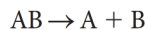
Most chemical reactions exhibit one of three recognizable patterns: They are either *synthesis*, *decomposition*, or *exchange reactions*.

When atoms or molecules combine to form a larger, more complex molecule, the process is a **synthesis**, or **combination, reaction**. A synthesis reaction always involves bond formation. It can be represented (using arbitrary letters) as



Synthesis reactions are the basis of constructive, or **anabolic**, activities in body cells, such as joining small molecules called amino acids into large protein molecules (Figure 2.11a). Synthesis reactions are conspicuous in rapidly growing tissues.

A **decomposition reaction** occurs when a molecule is broken down into smaller molecules or its constituent atoms:



Essentially, decomposition reactions are reverse synthesis reactions: Bonds are broken. Decomposition reactions underlie all degradative, or **catabolic**, processes in body cells. For example,

the bonds of glycogen molecules are broken to release simpler molecules of glucose sugar (Figure 2.11b).

Exchange, or **displacement, reactions** involve both synthesis and decomposition. Bonds are both made and broken. In an exchange reaction, parts of the reactant molecules change partners, so to speak, producing different product molecules:



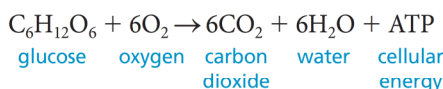
An exchange reaction occurs when ATP reacts with glucose and transfers its end phosphate group (indicated by a circled P in Figure 2.11c) to glucose, forming glucose-phosphate. At the same time, the ATP becomes ADP. This important reaction occurs whenever glucose enters a body cell, and it effectively traps the glucose fuel molecule inside the cell.

Another group of important chemical reactions in living systems is **oxidation-reduction reactions**, called **redox reactions** for short. Oxidation-reduction reactions are decomposition reactions in that they are the basis of all reactions in which food fuels are broken down for energy (that is, in which ATP is produced). They are also a special type of exchange reaction because electrons are exchanged between the reactants. The reactant losing the electrons is referred to as the *electron donor* and is said to be **oxidized**. The reactant taking up the transferred electrons is called the *electron acceptor* and is said to become **reduced**.

Redox reactions also occur when ionic compounds are formed. Recall that in the formation of NaCl (see Figure 2.6), sodium loses an electron to chlorine. Consequently, sodium is oxidized and becomes a sodium ion, and chlorine is reduced and becomes a chloride ion. However, not all oxidation-reduction

reactions involve *complete transfer* of electrons—some simply change the pattern of electron sharing in covalent bonds. For example, a substance is oxidized both by losing hydrogen atoms and by combining with oxygen. The common factor in these events is that electrons that formerly “belonged” to the reactant molecule are lost. The electrons are lost either entirely (as when hydrogen is removed and takes its electron with it) or relatively (as the shared electrons spend more time in the vicinity of the very electronegative oxygen atom).

To understand the importance of oxidation-reduction reactions in living systems, take a look at the overall equation for *cellular respiration*, which represents the major pathway by which glucose is broken down for energy in body cells:



As you can see, it is an oxidation-reduction reaction. Consider what happens to the hydrogen atoms (and their electrons). Glucose is oxidized to carbon dioxide as it loses hydrogen atoms, and oxygen is reduced to water as it accepts the hydrogen atoms. This reaction is described in detail in Chapter 24, along with other topics of cellular metabolism.

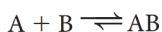
Energy Flow in Chemical Reactions

Because all chemical bonds represent stored chemical energy, all chemical reactions ultimately result in net absorption or release of energy. Reactions that release energy are called **exergonic reactions**. These reactions yield products with less energy than the initial reactants, along with energy that can be harvested for other uses. With a few exceptions, catabolic and oxidative reactions are exergonic.

In contrast, the products of energy-absorbing, or **endergonic**, reactions contain more potential energy in their chemical bonds than did the reactants. Anabolic reactions are typically energy-absorbing endergonic reactions. Essentially this is a case of “one hand washing the other”—the energy released when fuel molecules are broken down (oxidized) is captured in ATP molecules and then used to synthesize the complex biological molecules the body needs to sustain life.

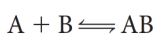
Reversibility of Chemical Reactions

All chemical reactions are theoretically reversible. If chemical bonds can be made, they can be broken, and vice versa. Reversibility is indicated by a double arrow. When the arrows differ in length, the longer arrow indicates the major direction in which the reaction proceeds:



In this example, the forward reaction (reaction going to the right) predominates. Over time, the product (AB) accumulates and the reactants (A and B) decrease in amount.

When the arrows are of equal length, as in



neither the forward reaction nor the reverse reaction is dominant. In other words, for each molecule of product (AB) formed, one product molecule breaks down, releasing the reactants A and B. Such a chemical reaction is said to be in a state of **chemical equilibrium**.

Once chemical equilibrium is reached, there is no further *net change* in the amounts of reactants and products unless more of either are added to the mix. Product molecules are still formed and broken down, but the balance established when equilibrium was reached (such as greater numbers of product molecules) remains unchanged.

Chemical equilibrium is analogous to the admission scheme used by many large museums in which tickets are sold according to time of entry. If 300 tickets are issued for the 9 AM admission, 300 people will be admitted when the doors open. Thereafter, when 6 people leave, 6 are admitted, and when another 15 people leave, 15 more are allowed in. There is a continual turnover, but the museum contains about 300 art lovers throughout the day.

All chemical reactions are reversible, but many biological reactions show so little tendency to go in the reverse direction that they are irreversible for all practical purposes. Chemical reactions that release energy will not go in the opposite direction unless energy is put back into the system. For example, when our cells break down glucose via the reactions of cellular respiration to yield carbon dioxide and water, some of the energy released is trapped in the bonds of ATP. Because the cells then use ATP's energy for various functions (and more glucose will be along with the next meal), this particular reaction is never reversed in our cells. Furthermore, if a product of a reaction is continuously removed from the reaction site, it is unavailable to take part in the reverse reaction. This situation occurs when the carbon dioxide that is released during glucose breakdown leaves the cell, enters the blood, and is eventually removed from the body by the lungs.

Factors Influencing the Rate of Chemical Reactions

What influences how quickly chemical reactions go? For atoms and molecules to react chemically in the first place, they must *collide* with enough force to overcome the repulsion between their electrons. Interactions between valence shell electrons—the basis of bond making and breaking—cannot occur long distance. The force of collisions depends on how fast the particles are moving. Solid, forceful collisions between rapidly moving particles in which valence shells overlap are much more likely to cause reactions than are collisions in which the particles graze each other lightly.

Temperature Increasing the temperature of a substance increases the kinetic energy of its particles and the force of their collisions. For this reason, chemical reactions proceed more quickly at higher temperatures.

Concentration Chemical reactions progress most rapidly when the reacting particles are present in high numbers, because the chance of successful collisions is greater. As the concentration of the reactants declines, the reaction slows. Chemical

equilibrium eventually occurs unless additional reactants are added or products are removed from the reaction site.

Particle Size Smaller particles move faster than larger ones (at the same temperature) and tend to collide more frequently and more forcefully. Hence, the smaller the reacting particles, the faster a chemical reaction goes at a given temperature and concentration.

Catalysts Many chemical reactions in nonliving systems can be speeded up simply by heating, but drastic increases in body temperature are life threatening because important biological molecules are destroyed. Still, at normal body temperatures, most chemical reactions would proceed far too slowly to maintain life were it not for the presence of catalysts. **Catalysts** (kat'ah-lists) are substances that increase the rate of chemical reactions without themselves becoming chemically changed or part of the product. Biological catalysts are called *enzymes* (en'zimz). Later in this chapter we describe how enzymes work.

✓ Check Your Understanding

13. Which reaction type—synthesis, decomposition, or exchange—occurs when fats are digested in your small intestine?
14. Why are many reactions that occur in living systems irreversible for all intents and purposes?
15. What specific name is given to decomposition reactions in which food fuels are broken down for energy?

For answers, see Appendix H.

PART 2 Biochemistry

Biochemistry is the study of the chemical composition and reactions of living matter. All chemicals in the body fall into one of two major classes: organic or inorganic compounds. **Organic compounds** contain carbon. All organic compounds are covalently bonded molecules, and many are large.

All other chemicals in the body are considered **inorganic compounds**. These include water, salts, and many acids and bases. Organic and inorganic compounds are equally essential for life. Trying to decide which is more valuable is like trying to decide whether the ignition system or the engine is more essential to run your car!

Inorganic Compounds

- ✓ Explain the importance of water and salts to body homeostasis.
- ✓ Define acid and base, and explain the concept of pH.

Water

Water is the most abundant and important inorganic compound in living material. It makes up 60–80% of the volume of most living cells. What makes water so vital to life? The answer lies in several properties:

- **High heat capacity.** Water has a high heat capacity. In other words, it absorbs and releases large amounts of heat before changing appreciably in temperature itself. This property of water prevents sudden changes in temperature caused by external factors, such as sun or wind exposure, or by internal conditions that release heat rapidly, such as vigorous muscle activity. As part of blood, water redistributes heat among body tissues, ensuring temperature homeostasis.
- **High heat of vaporization.** When water evaporates, or vaporizes, it changes from a liquid to a gas (water vapor). This transformation requires that large amounts of heat be absorbed to break the hydrogen bonds that hold water molecules together. This property is extremely beneficial when we sweat. As perspiration (mostly water) evaporates from our skin, large amounts of heat are removed from the body, providing efficient cooling.
- **Polar solvent properties.** Water is an unparalleled solvent. Indeed, it is often called the **universal solvent**. Biochemistry is “wet chemistry.” Biological molecules do not react chemically unless they are in solution, and virtually all chemical reactions occurring in the body depend on water’s solvent properties.

Because water molecules are polar, they orient themselves with their slightly negative ends toward the positive ends of the solutes, and vice versa, first attracting the solute molecules, and then surrounding them. This polarity of water explains why ionic compounds and other small reactive molecules (such as acids and bases) *dissociate* in water, their ions separating from each other and becoming evenly scattered in the water, forming true solutions (**Figure 2.12**).

Water also forms layers of water molecules, called **hydration layers**, around large charged molecules such as proteins, shielding them from the effects of other charged substances in the vicinity and preventing them from settling out of solution. Such protein-water mixtures are *biological colloids*. Blood plasma and cerebrospinal fluid (which surrounds the brain and spinal cord) are examples of colloids.

Water is the body’s major transport medium because it is such an excellent solvent. Nutrients, respiratory gases, and metabolic wastes carried throughout the body are dissolved in blood plasma, and many metabolic wastes are excreted from the body in urine, another watery fluid. Specialized molecules that lubricate the body (e.g., mucus) also use water as their dissolving medium.

- **Reactivity.** Water is an important *reactant* in many chemical reactions. For example, foods are digested to their building blocks by adding a water molecule to each bond to be broken. Such decomposition reactions are more specifically called **hydrolysis reactions** (hi-drol'i-sis; “water splitting”). Conversely, when large carbohydrate or protein molecules are synthesized from smaller molecules, a water molecule is removed for every bond formed, a reaction called **dehydration synthesis**.
- **Cushioning.** By forming a resilient cushion around certain body organs, water helps protect them from physical trauma.

The cerebrospinal fluid surrounding the brain exemplifies water's cushioning role.

Salts

A **salt** is an ionic compound containing cations other than H^+ and anions other than the hydroxyl ion (OH^-). As already noted, when salts are dissolved in water, they dissociate into their component ions (Figure 2.12). For example, sodium sulfate (Na_2SO_4) dissociates into two Na^+ ions and one SO_4^{2-} ion. It dissociates easily because the ions are already formed. All that remains is for water to overcome the attraction between the oppositely charged ions.

All ions are **electrolytes** (e-lek'tro-lits), substances that conduct an electrical current in solution. (Note that groups of atoms that bear an overall charge, such as sulfate, are called *polyatomic ions*.)

Salts commonly found in the body include $NaCl$, $CaCO_3$ (calcium carbonate), and KCl (potassium chloride). However, the most plentiful salts are the calcium phosphates that make bones and teeth hard. In their ionized form, salts play vital roles in body function. For instance, the electrolyte properties of sodium and potassium ions are essential for nerve impulse transmission and muscle contraction. Ionic iron forms part of the hemoglobin molecules that transport oxygen within red blood cells, and zinc and copper ions are important to the activity of some enzymes. Other important functions of the elements found in body salts are summarized in Table 2.1 on p. 26.

Homeostatic Imbalance 2.1

Maintaining proper ionic balance in our body fluids is one of the most crucial homeostatic roles of the kidneys. When this balance is severely disturbed, virtually nothing in the body works. All the physiological activities listed above and thousands of others are disrupted and grind to a stop. +

Acids and Bases

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

Acids

Acids have a sour taste, can react with (dissolve) many metals, and “burn” a hole in your rug. But for our purposes the most useful definition of an acid is a substance that releases **hydrogen ions** (H^+) in detectable amounts. Because a hydrogen ion is just a hydrogen nucleus, or “naked” proton, acids are also defined as **proton donors**.

When acids dissolve in water, they release hydrogen ions (protons) and anions. It is the concentration of protons that determines the acidity of a solution. The anions have little or no effect on acidity. For example, hydrochloric acid (HCl), an acid produced by stomach cells that aids digestion, dissociates into a proton and a chloride ion:

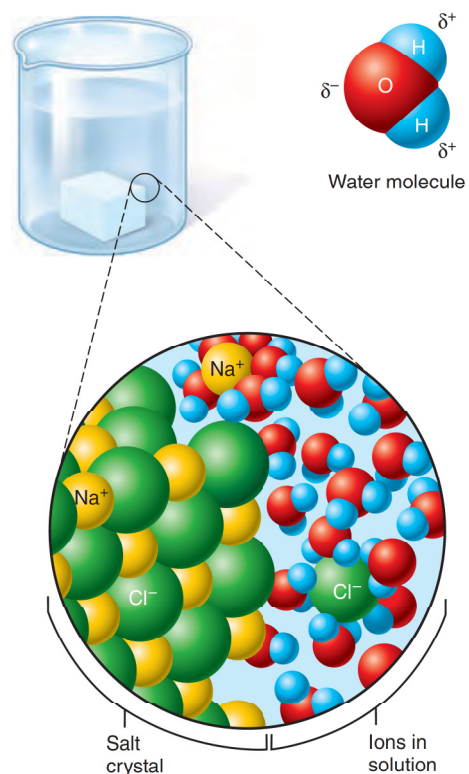
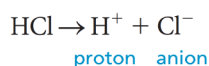
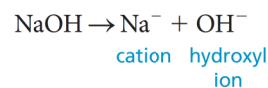


Figure 2.12 Dissociation of salt in water.

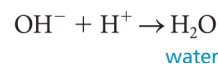
Other acids found or produced in the body include acetic acid ($HC_2H_3O_2$, commonly abbreviated as HAc), which is the acidic portion of vinegar; and carbonic acid (H_2CO_3). The molecular formula for an acid is easy to recognize because the hydrogen is written first.

Bases

Bases have a bitter taste, feel slippery, and are **proton acceptors**—that is, they take up hydrogen ions (H^+) in detectable amounts. Common inorganic bases include the *hydroxides* (hi-drok'sīds), such as magnesium hydroxide (milk of magnesia) and sodium hydroxide (lye). Like acids, hydroxides dissociate when dissolved in water, but in this case **hydroxyl ions** (OH^-) (hi-drok'sil) and cations are liberated. For example, ionization of sodium hydroxide ($NaOH$) produces a hydroxyl ion and a sodium ion, and the hydroxyl ion then binds to (accepts) a proton present in the solution. This reaction produces water and simultaneously reduces the acidity (hydrogen ion concentration) of the solution:

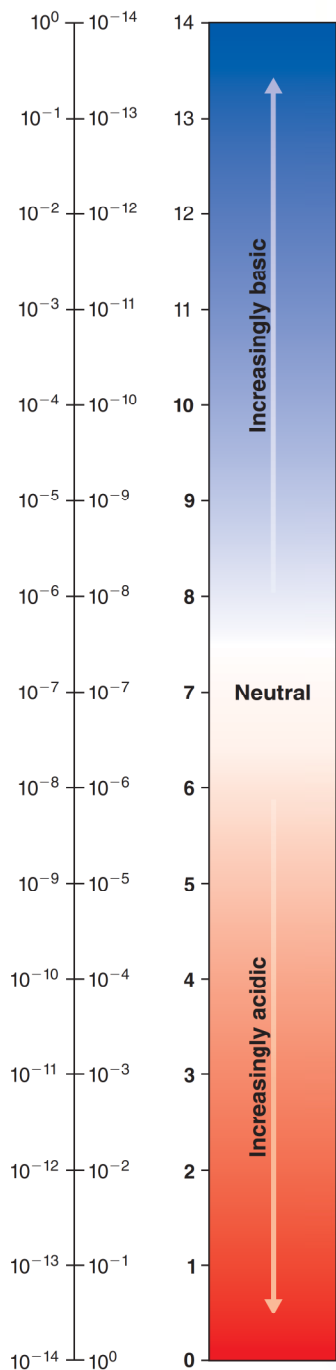


and then



Concentration (moles/liter)

[OH⁻] [H⁺] pH

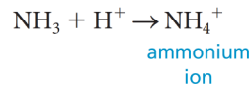


Examples

- 1M Sodium hydroxide (pH=14)
- Oven cleaner, lye (pH=13.5)
- Household ammonia (pH=10.5–11.5)
- Household bleach (pH=9.5)
- Egg white (pH=8)
- Blood (pH=7.4)
- Milk (pH=6.3–6.6)
- Black coffee (pH=5)
- Wine (pH=2.5–3.5)
- Lemon juice; gastric juice (pH=2)
- 1M Hydrochloric acid (pH=0)

Figure 2.13 The pH scale and pH values of representative substances. The pH scale is based on the number of hydrogen ions in solution. The actual concentrations of hydrogen ions, [H⁺], and hydroxyl ions, [OH⁻], in moles per liter are indicated for each pH value noted. At a pH of 7, [H⁺] = [OH⁻] and the solution is neutral.

Bicarbonate ion (HCO₃⁻), an important base in the body, is particularly abundant in blood. **Ammonia (NH₃)**, a common waste product of protein breakdown in the body, is also a base. It has one pair of unshared electrons that strongly attracts protons. By accepting a proton, ammonia becomes an ammonium ion:



pH: Acid-Base Concentration

The more hydrogen ions in a solution, the more acidic the solution is. Conversely, the greater the concentration of hydroxyl ions (the lower the concentration of H⁺), the more basic, or *alkaline* (al'kuh-lin), the solution becomes. The relative concentration of hydrogen ions in various body fluids is measured in concentration units called **pH units** (pe-äch').

The idea for a pH scale was devised by a Danish biochemist and part-time beer brewer named Sören Sørensen in 1909. He was searching for a convenient means of checking the acidity of his alcoholic product to prevent its spoilage by bacterial action. (Acidic conditions inhibit many bacteria.) The pH scale that resulted is based on the concentration of hydrogen ions in a solution, expressed in terms of moles per liter, or molarity. The pH scale runs from 0 to 14 and is *logarithmic*. In other words, each successive change of one pH unit represents a tenfold change in hydrogen ion concentration (Figure 2.13). The pH of a solution is thus defined as the negative logarithm of the hydrogen ion concentration [H⁺] in moles per liter, or -log[H⁺]. (Note that brackets [] indicate concentration of a substance.)

At a pH of 7 (at which [H⁺] is 10⁻⁷ M), the solution is *neutral*—neither acidic nor basic. The number of hydrogen ions exactly equals the number of hydroxyl ions (pH = pOH). Absolutely pure (distilled) water has a pH of 7.

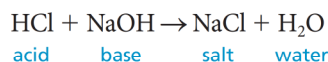
Solutions with a pH below 7 are acidic—the hydrogen ions outnumber the hydroxyl ions. The lower the pH, the more acidic the solution. A solution with a pH of 6 has ten times as many hydrogen ions as a solution with a pH of 7.

Solutions with a pH higher than 7 are alkaline, and the relative concentration of hydrogen ions decreases by a factor of 10 with each higher pH unit. Thus, solutions with pH values of 8 and 12 have, respectively, 1/10 and 1/100,000 (1/10 × 1/10 × 1/10 × 1/10 × 1/10 × 1/10) as many hydrogen ions as a solution of pH 7.

The approximate pH of several body fluids and of a number of common substances appears in Figure 2.13. Notice that as the hydrogen ion concentration decreases, the hydroxyl ion concentration rises, and vice versa.

Neutralization

What happens when acids and bases are mixed? They react with each other in displacement reactions to form water and a salt. For example, when hydrochloric acid and sodium hydroxide interact, sodium chloride (a salt) and water are formed.



This type of reaction is called a **neutralization reaction**, because the joining of H^+ and OH^- to form water neutralizes the solution. Although the salt produced is written in molecular form (NaCl), remember that it actually exists as dissociated sodium and chloride ions when dissolved in water.

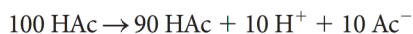
Buffers

Living cells are extraordinarily sensitive to even slight changes in the pH of the environment. In high concentrations, acids and bases are extremely damaging to living tissue. Imagine what would happen to all those hydrogen bonds in biological molecules with large numbers of free H^+ running around. (Can't you just hear those molecules saying "Why share hydrogen when I can have my own?")

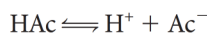
Homeostasis of acid-base balance is carefully regulated by the kidneys and lungs and by chemical systems (proteins and other types of molecules) called **buffers**. Buffers resist abrupt and large swings in the pH of body fluids by releasing hydrogen ions (acting as acids) when the pH begins to rise and by binding hydrogen ions (acting as bases) when the pH drops. Because blood comes into close contact with nearly every body cell, regulating its pH is particularly critical. Normally, blood pH varies within a very narrow range (7.35 to 7.45). If blood pH varies from these limits by more than a few tenths of a unit, it may be fatal.

To comprehend how chemical buffer systems operate, you must thoroughly understand strong and weak acids and bases. The first important concept is that the acidity of a solution reflects *only* the free hydrogen ions, not those still bound to anions. Consequently, acids that dissociate completely and irreversibly in water are called **strong acids**, because they can dramatically change the pH of a solution. Examples are hydrochloric acid and sulfuric acid. If we could count out 100 hydrochloric acid molecules and place them in 1 milliliter (ml) of water, we could expect to end up with 100 H^+ , 100 Cl^- , and no undissociated hydrochloric acid molecules in that solution.

Acids that do not dissociate completely, like carbonic acid (H_2CO_3) and acetic acid (HAc), are **weak acids**. If we were to place 100 acetic acid molecules in 1 ml of water, the reaction would be something like this:



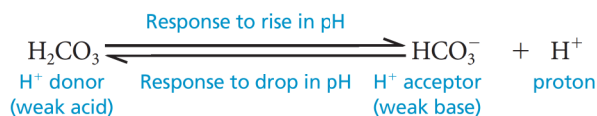
Because undissociated acids do not affect pH, the acetic acid solution is much less acidic than the HCl solution. Weak acids dissociate in a predictable way, and molecules of the intact acid are in dynamic equilibrium with the dissociated ions. Consequently, the dissociation of acetic acid may also be written as



This viewpoint allows us to see that if H^+ (released by a strong acid) is added to the acetic acid solution, the equilibrium will shift to the left and some H^+ and Ac^- will recombine to form HAc. On the other hand, if a strong base is added and the pH begins to rise, the equilibrium shifts to the right and more HAc molecules dissociate to release H^+ . This characteristic of weak acids allows them to play important roles in the chemical buffer systems of the body.

The concept of strong and weak bases is more easily explained. Remember that bases are proton acceptors. Thus, **strong bases** are those, like hydroxides, that dissociate easily in water and quickly tie up H^+ . On the other hand, sodium bicarbonate (commonly known as baking soda) ionizes incompletely and reversibly. Because it accepts relatively few protons, its released bicarbonate ion is considered a **weak base**.

Now let's examine how one buffer system helps to maintain pH homeostasis of the blood. Although there are other chemical blood buffers, the **carbonic acid–bicarbonate system** is a major one. Carbonic acid (H_2CO_3) dissociates reversibly, releasing bicarbonate ions (HCO_3^-) and protons (H^+):



The chemical equilibrium between carbonic acid (a weak acid) and bicarbonate ion (a weak base) resists changes in blood pH by shifting to the right or left as H^+ ions are added to or removed from the blood. As blood pH rises (becomes more alkaline due to the addition of a strong base), the equilibrium shifts to the right, forcing more carbonic acid to dissociate. Similarly, as blood pH begins to drop (becomes more acidic due to the addition of a strong acid), the equilibrium shifts to the left as more bicarbonate ions begin to bind with protons. As you can see, strong bases are replaced by a weak base (bicarbonate ion) and protons released by strong acids are tied up in a weak one (carbonic acid). In either case, the blood pH changes much less than it would in the absence of the buffering system. We discuss acid-base balance and buffers in more detail in Chapter 26.

Check Your Understanding

- Water makes up 60–80% of living matter. What property makes it an excellent solvent?
- Salts are electrolytes. What does that mean?
- Which ion is responsible for increased acidity?
- To minimize the sharp pH shift that occurs when a strong acid is added to a solution, is it better to add a weak base or a strong base? Why?

For answers, see Appendix H.

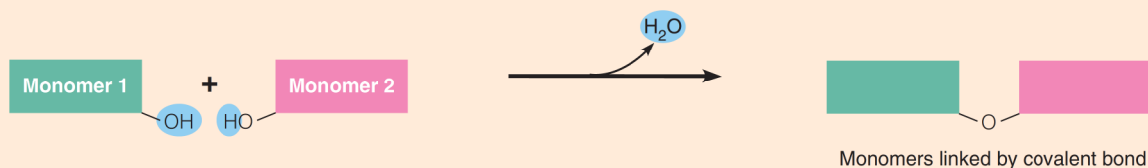
Organic Compounds

- Describe and compare the building blocks, general structures, and biological functions of carbohydrates and lipids.
- Explain the role of dehydration synthesis and hydrolysis in forming and breaking down organic molecules.

Molecules unique to living systems—carbohydrates, lipids (fats), proteins, and nucleic acids—all contain carbon and hence are organic compounds. Organic compounds are generally distinguished by the fact that they contain carbon, and inorganic

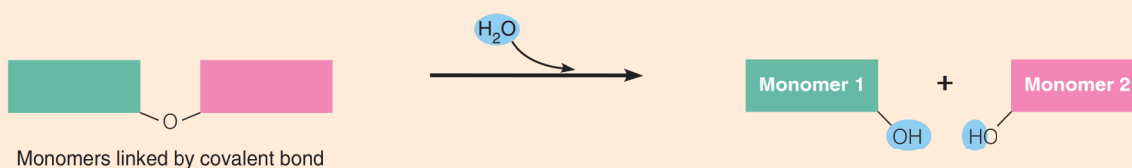
(a) Dehydration synthesis

Monomers are joined by removal of OH from one monomer and removal of H from the other at the site of bond formation.



(b) Hydrolysis

Monomers are released by the addition of a water molecule, adding OH to one monomer and H to the other.



(c) Example reactions

Dehydration synthesis of sucrose and its breakdown by hydrolysis

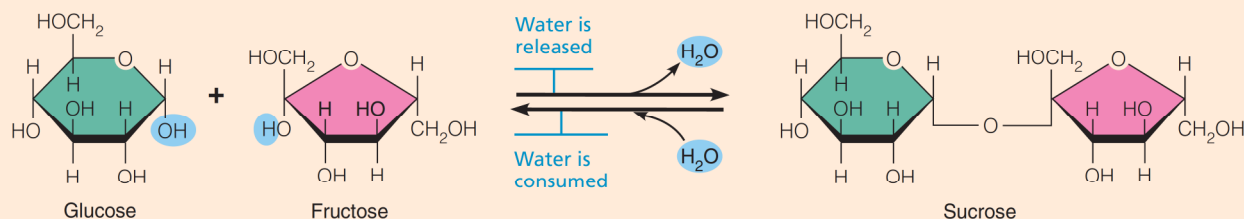


Figure 2.14 Dehydration synthesis and hydrolysis. Biological molecules are formed from their monomers, or units, by dehydration synthesis and broken down to the monomers by hydrolysis reactions.

compounds are defined as compounds that lack carbon. You should be aware of a few irrational exceptions to this generalization: Carbon dioxide and carbon monoxide, for example, contain carbon but are considered inorganic compounds.

For the most part, organic molecules 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). The most important functional groups involved in biochemical reactions are illustrated in Appendix B.

What makes carbon so special that “living” chemistry depends on its presence? To begin with, no other *small* atom is so precisely **electroneutral**. The consequence of its electroneutrality is that carbon never loses or gains electrons. Instead, it always shares them. Furthermore, with four valence shell electrons, carbon forms four covalent bonds with other

elements, as well as with other carbon atoms. As a result, carbon can help form long, chainlike molecules (common in fats), ring structures (typical of carbohydrates and steroids), and many other structures that are uniquely suited for specific roles in the body.

As you will see shortly, many biological molecules (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 2.14**). During dehydration synthesis, a hydrogen atom is removed from one monomer and a hydroxyl group is removed from the monomer it is to be joined with. As a covalent bond unites the monomers, a water molecule is released. This removal of a water molecule at the bond site occurs each time a monomer is added to the growing polymer chain.

Carbohydrates

Carbohydrates, a group of molecules that includes sugars and starches, represent 1–2% of cell mass. Carbohydrates contain carbon, hydrogen, and oxygen, and generally the hydrogen and oxygen atoms occur in the same 2:1 ratio as in water. This ratio is reflected in the word *carbohydrate* (“hydrated carbon”).

A carbohydrate can be classified according to size and solubility as a monosaccharide (“one sugar”), disaccharide (“two sugars”), or polysaccharide (“many sugars”). Monosaccharides are the monomers, or building blocks, of the other carbohydrates. In general, the larger the carbohydrate molecule, the less soluble it is in water.

Monosaccharides

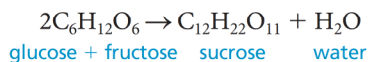
Monosaccharides (mon’o-sak’ah-rīdz), or *simple sugars*, are single-chain or single-ring structures containing from three to seven carbon atoms (Figure 2.15a). Usually the carbon, hydrogen, and oxygen atoms occur in the ratio 1:2:1, so a general formula for a monosaccharide is $(\text{CH}_2\text{O})_n$, where n is the number of carbons in the sugar. Glucose, for example, has six carbon atoms, and its molecular formula is $\text{C}_6\text{H}_{12}\text{O}_6$. Ribose, with five carbons, is $\text{C}_5\text{H}_{10}\text{O}_5$.

Monosaccharides are named generically according to the number of carbon atoms they contain. Most important in the body are the pentose (five-carbon) and hexose (six-carbon) sugars. For example, the pentose *deoxyribose* (de-ok’si-ri’bōs) is part of DNA, and *glucose*, a hexose, is blood sugar.

Two other hexoses, *galactose* and *fructose*, are **isomers** (i’so-mers) of glucose. That is, they have the same molecular formula ($\text{C}_6\text{H}_{12}\text{O}_6$), but as you can see in Figure 2.15a, their atoms are arranged differently, giving them different chemical properties.

Disaccharides

A **disaccharide** (di-sak’ah-rīd), or *double sugar*, is formed when two monosaccharides are joined by *dehydration synthesis* (Figure 2.14a, c). In this synthesis reaction, a water molecule is lost as the bond is made, as illustrated by the synthesis of sucrose (soo’krōs):



Notice that the molecular formula for sucrose contains two hydrogen atoms and one oxygen atom less than the total number of hydrogen and oxygen atoms in glucose and fructose, because a water molecule is released during bond formation.

Important disaccharides in the diet are *sucrose* (glucose + fructose), which is cane or table sugar; *lactose* (glucose + galactose), found in milk; and *maltose* (glucose + glucose), also called malt sugar (Figure 2.15b). Disaccharides are too large to pass through cell membranes, so they must be digested to their simple sugar units to be absorbed from the digestive tract into the blood. This decomposition process is *hydrolysis*, essentially the reverse of dehydration synthesis (Figure 2.14a, b). A water

molecule is added to each bond, breaking the bonds and releasing the simple sugar units.

Polysaccharides

Polysaccharides (pol’e-sak’ah-rīdz) are polymers of simple sugars linked together by dehydration synthesis. Because polysaccharides are large, fairly insoluble molecules, they are ideal storage products. Another consequence of their large size is that they lack the sweetness of the simple and double sugars.

Only two polysaccharides are of major importance to the body: starch and glycogen. Both are polymers of glucose. Only their degree of branching differs.

Starch is the storage carbohydrate formed by plants. The number of glucose units composing a starch molecule is high and variable. When we eat starchy foods such as grain products and potatoes, the starch must be digested for its glucose units to be absorbed. We are unable to digest *cellulose*, another polysaccharide found in all plant products. However, it is important in providing the *bulk* (one form of fiber) that helps move feces through the colon.

Glycogen (gli’ko-jen), the storage carbohydrate of animal tissues, is stored primarily in skeletal muscle and liver cells. Like starch, it is highly branched and is a very large molecule (Figure 2.15c). When blood sugar levels drop sharply, liver cells break down glycogen and release its glucose units to the blood. Since there are many branch endings from which glucose can be released simultaneously, body cells have almost instant access to glucose fuel.

Carbohydrate Functions

The major function of carbohydrates in the body is to provide a ready, easily used source of cellular fuel. Most cells can use only a few types of simple sugars, and glucose is at the top of the “cellular menu.” As described in our earlier discussion of oxidation-reduction reactions (pp. 36–37), glucose is broken down and oxidized within cells. During these chemical reactions, electrons are transferred. This relocation of electrons releases the bond energy stored in glucose, and this energy is used to synthesize ATP. When ATP supplies are sufficient, 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 personal experience with this conversion process!

Only small amounts of carbohydrates are used for structural purposes. For example, some sugars are found in our genes. Others are attached to the external surfaces of cells where they act as “road signs” to guide cellular interactions.

Lipids

Lipids are insoluble in water but dissolve readily in other lipids and in organic solvents such as alcohol and ether. Like carbohydrates, all lipids contain carbon, hydrogen, and oxygen, but the proportion of oxygen in lipids is much lower. In addition, phosphorus is found in some of the more complex lipids. Lipids include *triglycerides*, *phospholipids* (fos’fo-lip’idz), *steroids* (stē’roidz), and a number of other lipid substances. Table 2.2 on p. 46 gives the locations and functions of some lipids found in the body.

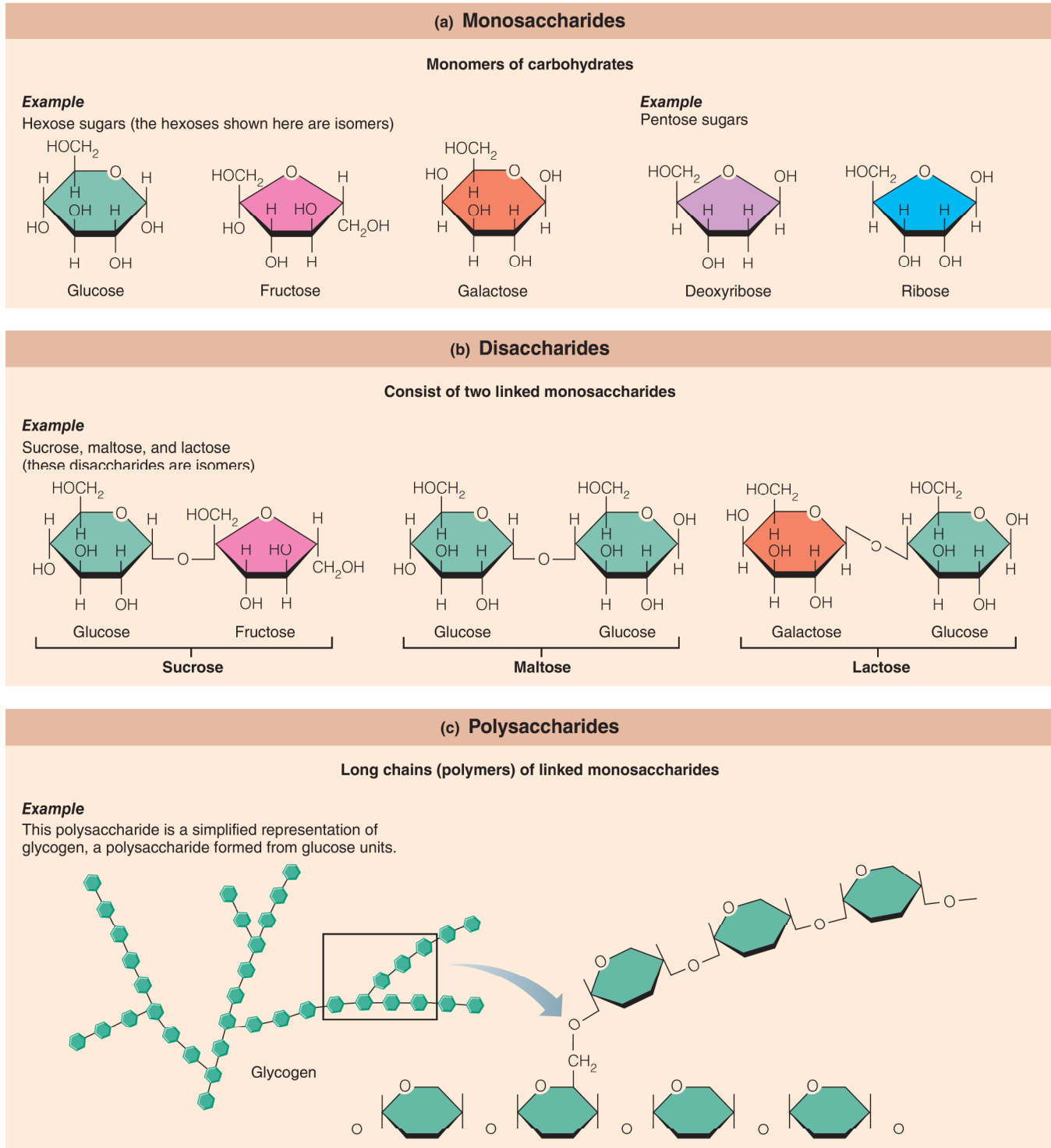
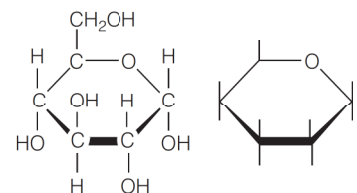


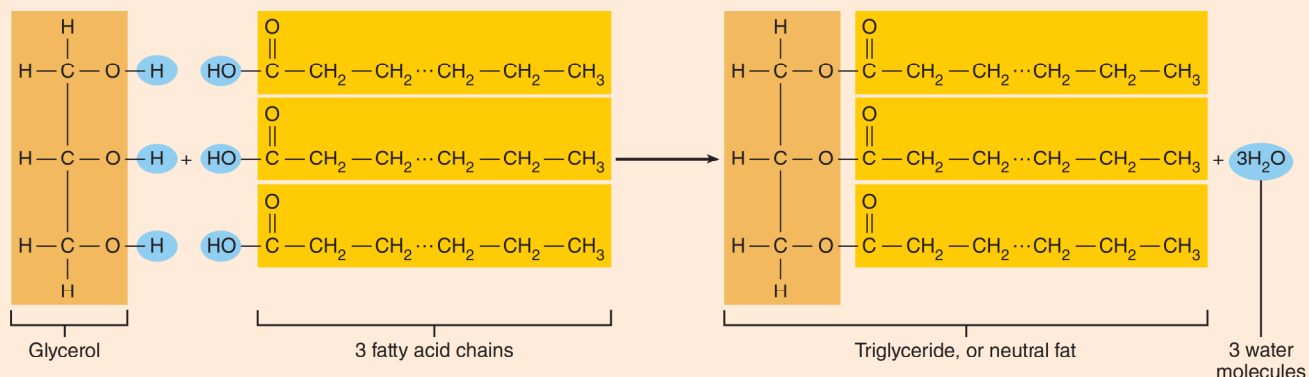
Figure 2.15 Carbohydrate molecules important to the body.*

*Notice that in Figure 2.15 the carbon (C) atoms present at the angles of the carbohydrate ring structures are not illustrated and in Figure 2.15c only the oxygen atoms and one CH₂ group are shown. The illustrations at right give an example of this shorthand style: The full structure of glucose is on the left and the shorthand structure on the right. This style is used for nearly all organic ringlike structures illustrated in this chapter.



(a) Triglyceride formation

Three fatty acid chains are bound to glycerol by dehydration synthesis.

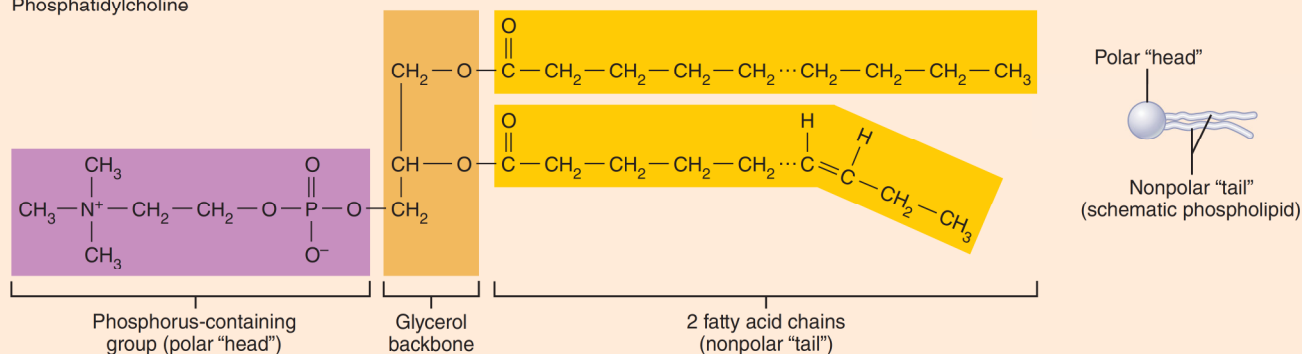


(b) "Typical" structure of a phospholipid molecule

Two fatty acid chains and a phosphorus-containing group are attached to the glycerol backbone.

Example

Phosphatidylcholine

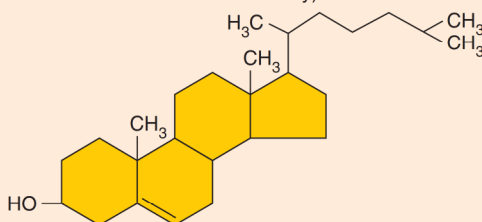


(c) Simplified structure of a steroid

Four interlocking hydrocarbon rings form a steroid.

Example

Cholesterol (cholesterol is the basis for all steroids formed in the body)

**Triglycerides (Neutral Fats)**

Triglycerides (tri-glis'er-idz), also called **neutral fats**, are commonly known as *fats* when solid or *oils* when liquid. A triglyceride is composed of two types of building blocks, **fatty acids**

Figure 2.16 Lipids. The general structure of (a) triglycerides, or neutral fats, (b) phospholipids, and (c) cholesterol.

and **glycerol** (glis'er-ol), in a 3:1 ratio of fatty acids to glycerol (**Figure 2.16a**). Fatty acids are linear chains of carbon and hydrogen atoms (hydrocarbon chains) with an organic acid group (—COOH) at one end. Glycerol is a modified simple sugar (a sugar alcohol).

Fat synthesis involves attaching three fatty acid chains to a single glycerol molecule by dehydration synthesis. The result is an E-shaped molecule. The glycerol backbone is the same in all triglycerides, but the fatty acid chains vary, resulting in different kinds of fats and oils.

These are large molecules, often consisting of hundreds of atoms, and ingested fats and oils must be broken down to their building blocks before they can be absorbed. Their hydrocarbon chains make the triglycerides nonpolar molecules. Because polar and nonpolar molecules do not interact, oil (or fats) and water do not mix. Consequently, triglycerides provide the body's

Table 2.2 Representative Lipids Found in the Body

LIPID TYPE	LOCATION/FUNCTION
Triglycerides (Neutral Fats)	
	Fat deposits (in subcutaneous tissue and around organs) protect and insulate body organs, and are the major source of <i>stored</i> energy in the body.
Phospholipids (phosphatidylcholine; cephalin; others)	
	Chief components of cell membranes. Participate in the transport of lipids in plasma. Prevalent in nervous tissue.
Steroids	
Cholesterol	The structural basis for manufacture of all body steroids. A component of cell membranes.
Bile salts	These breakdown products of cholesterol are released by the liver into the digestive tract, where they aid fat digestion and absorption.
Vitamin D	A fat-soluble vitamin produced in the skin on exposure to UV radiation. Necessary for normal bone growth and function.
Sex hormones	Estrogen and progesterone (female hormones) and testosterone (a male hormone) are produced in the gonads. Necessary for normal reproductive function.
Adrenocortical hormones	Cortisol, a glucocorticoid, is a metabolic hormone necessary for maintaining normal blood glucose levels. Aldosterone helps to regulate salt and water balance of the body by targeting the kidneys.
Other Lipoid Substances	
Fat-soluble vitamins:	
A	Ingested in orange-pigmented vegetables and fruits. Converted in the retina to retinal, a part of the photoreceptor pigment involved in vision.
E	Ingested in plant products such as wheat germ and green leafy vegetables. Claims have been made (but not proved in humans) that it promotes wound healing, contributes to fertility, and may help to neutralize highly reactive particles called free radicals believed to be involved in triggering some types of cancer.
K	Made available to humans largely by the action of intestinal bacteria. Also prevalent in a wide variety of foods. Necessary for proper clotting of blood.
Eicosanoids (prostaglandins; leukotrienes; thromboxanes)	Group of molecules derived from fatty acids found in all cell membranes. The potent prostaglandins have diverse effects, including stimulation of uterine contractions, regulation of blood pressure, control of gastrointestinal tract motility, and secretory activity. Both prostaglandins and leukotrienes are involved in inflammation. Thromboxanes are powerful vasoconstrictors.
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).

most efficient and compact form of stored energy, and when they are oxidized, they yield large amounts of energy.

Triglycerides are found mainly beneath the skin, where they insulate the deeper body tissues from heat loss and protect them from mechanical trauma. For example, women are usually more successful English Channel swimmers than men. Their success is due partly to their thicker subcutaneous fatty layer, which helps insulate them from the bitterly cold water of the Channel.

The length of a triglyceride's fatty acid chains and their degree of *saturation* with H atoms 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**. Their fatty acid chains are straight and, at room temperature, the molecules of a saturated fat are packed closely together, forming

a solid. Fatty acids that contain one or more double bonds between carbon atoms are said to be **unsaturated** (**monounsaturated** and **polyunsaturated**, respectively). The double bonds cause the fatty acid chains to kink so that they cannot be packed closely enough to solidify. Hence, triglycerides with short fatty acid chains or unsaturated fatty acids are oils (liquid at room temperature) and are typical of plant lipids. Examples include olive and peanut oils (rich in monounsaturated fats) and corn, soybean, and safflower oils, which contain a high percentage of 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 said to be more “heart healthy.”

Trans fats, common in many margarines and baked products, are oils that have been solidified by addition of H atoms at sites of carbon double bonds. They have recently been branded as increasing the risk of heart disease even more than the solid animal fats. Conversely, the **omega-3 fatty acids**, found naturally in cold-water fish, appear to decrease the risk of heart disease and some inflammatory diseases.

Phospholipids

Phospholipids are modified triglycerides. Specifically, they are diglycerides with a phosphorus-containing group and two, rather than three, fatty acid chains (Figure 2.16b). The phosphorus-containing group gives phospholipids their distinctive chemical properties. Although the hydrocarbon portion (the “tail”) of the molecule is nonpolar and interacts only with nonpolar molecules, the phosphorus-containing part (the “head”) is polar and attracts other polar or charged particles, such as water or ions. This unique characteristic of phospholipids allows them to be used as the chief material for building cellular membranes. Some biologically important phospholipids and their functions are listed in Table 2.2.

Steroids

Structurally, steroids differ quite a bit from fats and oils. **Steroids** are basically flat molecules made of four interlocking hydrocarbon rings. Like triglycerides, steroids are fat soluble and contain little oxygen. The single most important molecule in our steroid chemistry is *cholesterol* (ko-les’ter-ol) (Figure 2.16c). We ingest cholesterol in animal products such as eggs, meat, and cheese, and our liver produces some.

Cholesterol has earned bad press because of its role in atherosclerosis, but it is essential for human life. Cholesterol is found in cell membranes and is the raw material for synthesis of vitamin D, steroid hormones, and bile salts. 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 a total lack of the corticosteroids produced by the adrenal glands is fatal.

Eicosanoids

The **eicosanoids** (i-ko’sah-noyds) are diverse lipids chiefly derived from a 20-carbon fatty acid (arachidonic acid) found in all cell membranes. Most important of these are the *prostaglandins* and their relatives, which play roles in various body processes including blood clotting, regulation of blood pressure, inflammation, and labor contractions (Table 2.2). Their synthesis and inflammatory actions are blocked by NSAIDs (nonsteroidal anti-inflammatory drugs) and the newer COX inhibitors.

✓ Check Your Understanding

20. What are the monomers of carbohydrates called? Which monomer is blood sugar?
21. What is the animal form of stored carbohydrate called?
22. How do triglycerides differ from phospholipids in body function and location?

23. What is the result of hydrolysis reactions and how are these reactions accomplished in the body?

For answers, see Appendix H.

Proteins

- ✓ Describe the four levels of protein structure.
- ✓ Indicate the function of molecular chaperones.
- ✓ Describe enzyme action.

The full set of proteins made by the body, called the *proteome*, and the way those proteins network in the body or change with disease, is a matter of intense biotech research.

Protein composes 10–30% of cell mass and is the basic structural material of the body. However, not all proteins are construction materials. Many play vital roles in cell function. Proteins, which include enzymes (biological catalysts), hemoglobin of the blood, and contractile proteins of muscle, have the most varied functions of any molecules in the body. All proteins contain carbon, oxygen, hydrogen, and nitrogen, and many contain sulfur as well.

Amino Acids and Peptide Bonds

The building blocks of proteins are molecules called **amino acids**, of which there are 20 common types (see Appendix C). All amino acids have two important functional groups: a basic group called an *amine* (ah’mēn) *group* ($-\text{NH}_2$), and an organic *acid group* ($-\text{COOH}$). An amino acid may therefore act either as a base (proton acceptor) or an acid (proton donor). All amino acids are identical except for a single group of atoms called their *R group*. Hence, it is differences in the R group that make each amino acid chemically unique, as the examples in Figure 2.17 show.

Proteins are long chains of amino acids joined together by dehydration synthesis, with the amine end of one amino acid linked to the acid end of the next. The resulting bond produces a characteristic arrangement of linked atoms called a **peptide bond** (Figure 2.18). Two united amino acids form a *dipeptide*, three a *tripeptide*, and ten or more a *polypeptide*. Although polypeptides containing more than 50 amino acids are called proteins, most proteins are **macromolecules**, large, complex molecules containing from 100 to over 10,000 amino acids.

Because each type of amino acid has distinct properties, the sequence in which they are bound together produces proteins that vary widely in both structure and function. We can think of the 20 amino acids as a 20-letter “alphabet” used in specific combinations to form “words” (proteins). Just as a change in one letter can produce a word with an entirely different meaning (flour → floor) or that is nonsensical (flour → flocl), changes in the kinds or positions of amino acids can yield proteins with different functions or proteins that are nonfunctional. Nevertheless, there are thousands of different proteins in the body, each with distinct functional properties, and all constructed from different combinations of the 20 common amino acids.

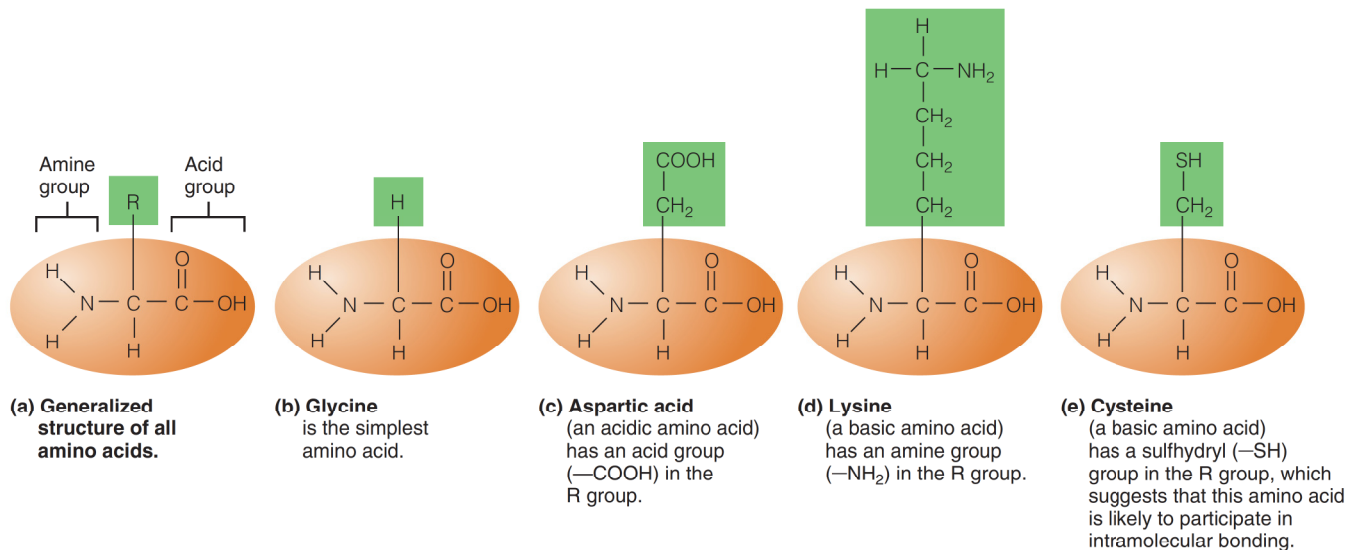


Figure 2.17 Amino acid structures. All amino acids have both an amine group ($-\text{NH}_2$) and an acid group ($-\text{COOH}$). They differ only in their R groups (green). Differences in their R groups allow them to function differently in the body.

Structural Levels of Proteins

Proteins can be described in terms of four structural levels. The linear sequence of amino acids composing the polypeptide chain is called the *primary structure* of a protein. This structure, which resembles a strand of amino acid “beads,” is the backbone of the protein molecule (**Figure 2.19a**).

Proteins do not normally exist as simple, linear chains of amino acids. Instead, they twist or bend upon themselves to form a more complex *secondary structure*. The most common type of secondary structure is the **alpha (α)-helix**, which resembles a Slinky toy or the coils of a telephone cord (**Figure 2.19b**). The α -helix is formed by coiling of the primary chain and is stabilized by hydrogen bonds formed between NH and CO groups in amino acids in the primary chain which are approximately

four amino acids apart. Hydrogen bonds in α -helices 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 an accordion’s bellows (**Figure 2.19b**). Notice that 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. A single polypeptide chain may exhibit both types of secondary structure at various places along its length.

Many proteins have *tertiary structure* (ter’she-a’re), the next higher level of complexity, which is superimposed on secondary structure. Tertiary structure is achieved when α -helical or

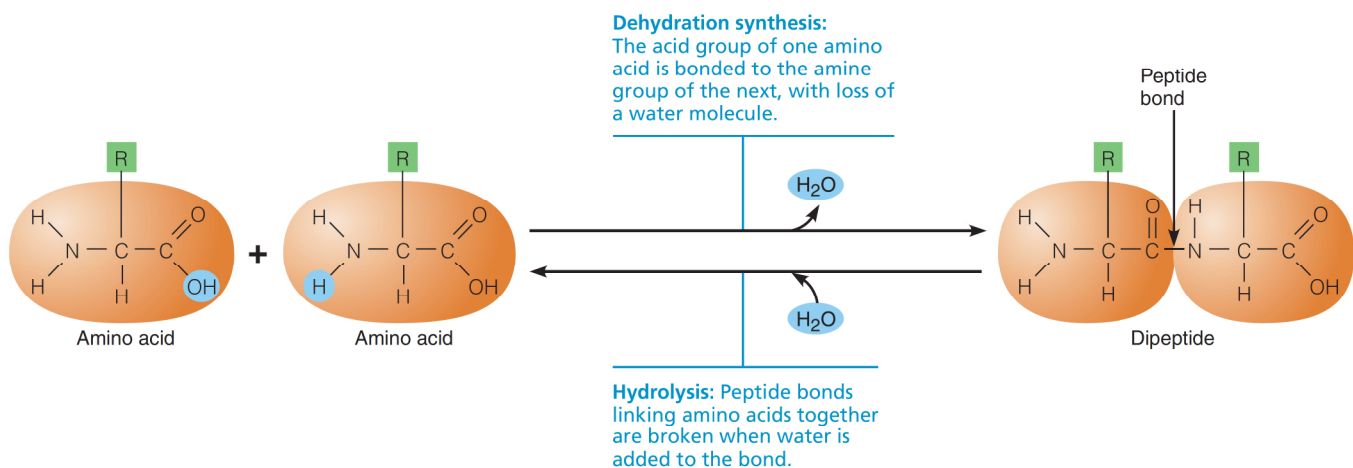


Figure 2.18 Amino acids are linked together by peptide bonds. Peptide bonds are formed by dehydration synthesis and broken by hydrolysis reactions.

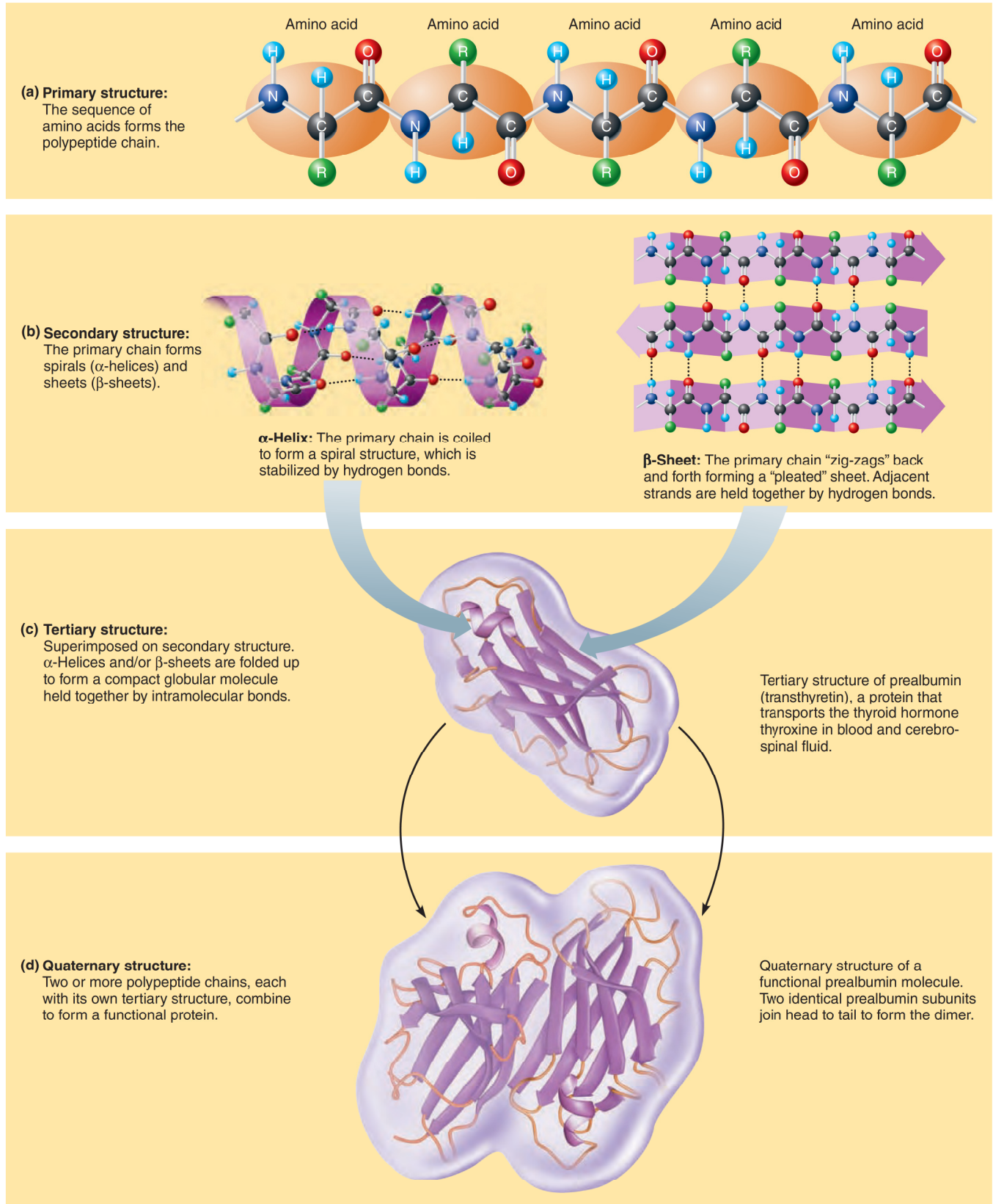


Figure 2.19 Levels of protein structure.

β -pleated regions of the polypeptide chain fold upon one another to produce a compact ball-like, or *globular*, molecule (Figure 2.19c). The unique structure is maintained by both covalent and hydrogen bonds between amino acids that are often far apart in the primary chain.

When two or more polypeptide chains aggregate in a regular manner to form a complex protein, the protein has *quaternary structure* (kwah'ter-na're). Prealbumin, a protein that transports thyroid hormone in the blood, exhibits this structural level (Figure 2.19d).

How do these different levels of structure arise? Although a protein with tertiary or quaternary structure looks a bit like a clump of congealed pasta, the ultimate overall structure of any protein is very specific and is dictated by its primary structure. In other words, the types and relative positions of amino acids in the protein backbone determine where bonds can form to produce the complex coiled or folded structures that keep water-loving amino acids near the surface and water-fleeing amino acids buried in the protein's core. In addition, cells decorate many proteins by attaching sugars or fatty acids to them in ways that are difficult to imagine or predict.

Fibrous and Globular Proteins

The overall structure of a protein determines its biological function. In general, proteins are classified according to their overall appearance and shape as either fibrous or globular.

Fibrous proteins are extended and strandlike. Some exhibit only secondary structure, but most have tertiary or even quaternary structure as well. For example, *collagen* (kol'ah-jen) is a composite of the helical tropocollagen molecules packed together side by side to form a strong ropelike structure. Fibrous proteins are insoluble in water, and very stable—qualities ideal for providing mechanical support and tensile strength to the body's tissues. Besides collagen, which is the single most abundant protein in the body, the fibrous proteins include keratin, elastin, and certain contractile proteins of muscle (Table 2.3). Because fibrous proteins are the chief building materials of the body, they are also known as **structural proteins**.

Globular proteins are compact, spherical proteins that have at least tertiary structure. Some also exhibit quaternary structure. The globular proteins are water-soluble, chemically active molecules, and they play crucial roles in virtually all biological processes. Consequently, this group is also called **functional proteins**. Some (antibodies) help to provide immunity, others (protein-based hormones) regulate growth and development, and still others (enzymes) are catalysts that oversee just about every chemical reaction in the body. The roles of these and selected other proteins found in the body are summarized in Table 2.3.

Protein Denaturation

Fibrous proteins are stable, but globular proteins are quite the opposite. The activity of a protein depends on its specific three-dimensional structure, and intramolecular bonds, particularly hydrogen bonds, are important in maintaining that structure. However, hydrogen bonds are fragile and easily broken by many

chemical and physical factors, such as excessive acidity or temperature. Although individual proteins vary in their sensitivity to environmental conditions, hydrogen bonds begin to break when the pH drops or the temperature rises above normal (physiological) levels, causing proteins to unfold and lose their specific three-dimensional shape. In this condition, a protein is said to be **denatured**.

Fortunately, the disruption is reversible in most cases, and the “scrambled” protein regains its native structure when desirable conditions are restored. However, if the temperature or pH change is so extreme that protein structure is damaged beyond repair, the protein is *irreversibly denatured*. The coagulation of egg white (primarily albumin protein) that occurs when you boil or fry an egg is an example of irreversible protein denaturation. There is no way to restore the white, rubbery protein to its original translucent form.

When globular proteins are denatured, they can no longer perform their physiological roles because their function depends on the presence of specific arrangements of atoms, called **active sites**, on their surfaces. The active sites are regions that fit and interact chemically with other molecules of complementary shape and charge. Because atoms contributing to an active site may actually be far apart in the primary chain, disruption of intramolecular bonds separates them and destroys the active site. For example, hemoglobin becomes totally unable to bind and transport oxygen when blood pH is too acidic, because the structure needed for its function has been destroyed.

We will describe most types of body proteins in conjunction with the organ systems or functional processes to which they are closely related. However, two groups of proteins—*molecular chaperones* and *enzymes*—are intimately involved in the normal functioning of all cells, so we will consider these incredibly complex molecules here.

✓ Check Your Understanding

24. What does the name “amino acid” tell you about the structure of this molecule?
25. What is the primary structure of proteins?
26. What are the two types of secondary structure in proteins?

For answers, see Appendix H.

Molecular Chaperones

In addition to enzymes, all cells contain a class of unrelated globular proteins called **molecular chaperones** which, among other things, help proteins to achieve their functional three-dimensional structure. Although its amino acid sequence determines the precise way a protein folds, the folding process also requires the help of molecular chaperones to ensure that the folding is quick and accurate. Molecular chaperones can associate with a broad range of “client” proteins, allowing them to perform a dizzying array of jobs. For example, specific molecular chaperones

- Prevent accidental, premature, or incorrect folding of polypeptide chains or their association with other polypeptides
- Aid the desired folding and association process

Table 2.3 Representative Types of Proteins in the Body

CLASSIFICATION ACCORDING TO		
OVERALL STRUCTURE	GENERAL FUNCTION	EXAMPLES FROM THE BODY
Fibrous		
	Structural framework/ mechanical support	<i>Collagen</i> , found in all connective tissues, is the single most abundant protein in the body. It is responsible for the tensile strength of bones, tendons, and ligaments. <i>Keratin</i> is the structural protein of hair and nails and a water-resistant material of skin. <i>Elastin</i> is found, along with collagen, where durability and flexibility are needed, such as in the ligaments that bind bones together. <i>Spectrin</i> internally reinforces and stabilizes the surface membrane of some cells, particularly red blood cells. <i>Dystrophin</i> reinforces and stabilizes the surface membrane of muscle cells. <i>Titin</i> helps organize the intracellular structure of muscle cells and accounts for the elasticity of skeletal muscles.
	Movement	<i>Actin</i> and <i>myosin</i> , contractile proteins, are found in substantial amounts in muscle cells, where they cause muscle cell shortening (contraction); they also function in cell division in all cell types. Actin is important in intracellular transport, particularly in nerve cells.
Globular		
	Catalysis	Protein enzymes are essential to virtually every biochemical reaction in the body; they increase the rates of chemical reactions by at least a millionfold. Examples include <i>salivary amylase</i> (in saliva), which catalyzes the breakdown of starch, and <i>oxidase enzymes</i> , which act to oxidize food fuels.
	Transport	<i>Hemoglobin</i> transports oxygen in blood, and <i>lipoproteins</i> transport lipids and cholesterol. Other transport proteins in the blood carry iron, hormones, or other substances. Some globular proteins in plasma membranes are involved in membrane transport (as carriers or channels).
	Regulation of pH	Many plasma proteins, such as <i>albumin</i> , function reversibly as acids or bases, thus acting as buffers to prevent wide swings in blood pH.
	Regulation of metabolism	<i>Peptide</i> and <i>protein hormones</i> help to regulate metabolic activity, growth, and development. For example, <i>growth hormone</i> is an anabolic hormone necessary for optimal growth; <i>insulin</i> helps regulate blood sugar levels.
	Body defense	<i>Antibodies</i> (immunoglobulins) are specialized proteins released by immune cells that recognize and inactivate foreign substances (bacteria, toxins, some viruses). <i>Complement proteins</i> , which circulate in blood, enhance both immune and inflammatory responses.
	Protein management	<i>Molecular chaperones</i> aid folding of new proteins in both healthy and damaged cells and transport of metal ions into and within the cell. They also promote breakdown of damaged proteins.

- Help to translocate proteins and certain metal ions (copper, iron, zinc) across cell membranes
- Promote the breakdown of damaged or denatured proteins
- Interact with immune cells to trigger the immune response to diseased cells in the body

The first such proteins discovered were called *heat shock proteins* (*hsp*) because they seemed to protect cells from the destructive effects of heat. It was later found that these proteins are produced in response to a variety of traumatizing stimuli—for example, in the oxygen-deprived cells of a heart attack patient—and the name *stress proteins* replaced *hsp* for that particular group of molecular

chaperones. It is now clear that these proteins are vitally important to cell function in all types of stressful circumstances. Although a lot of environmental stress can turn your hair gray or cause cardiac problems, it seems that a little stress can delay aging. As part of the heat shock response, special protein repair molecules patch up damaged proteins and refold them correctly, helping to hold off the grim reaper and prolong life.

Enzymes and Enzyme Activity

Enzymes are globular proteins that act as biological catalysts. *Catalysts* are substances that regulate and accelerate the rate of biochemical reactions but are not used up or changed in those

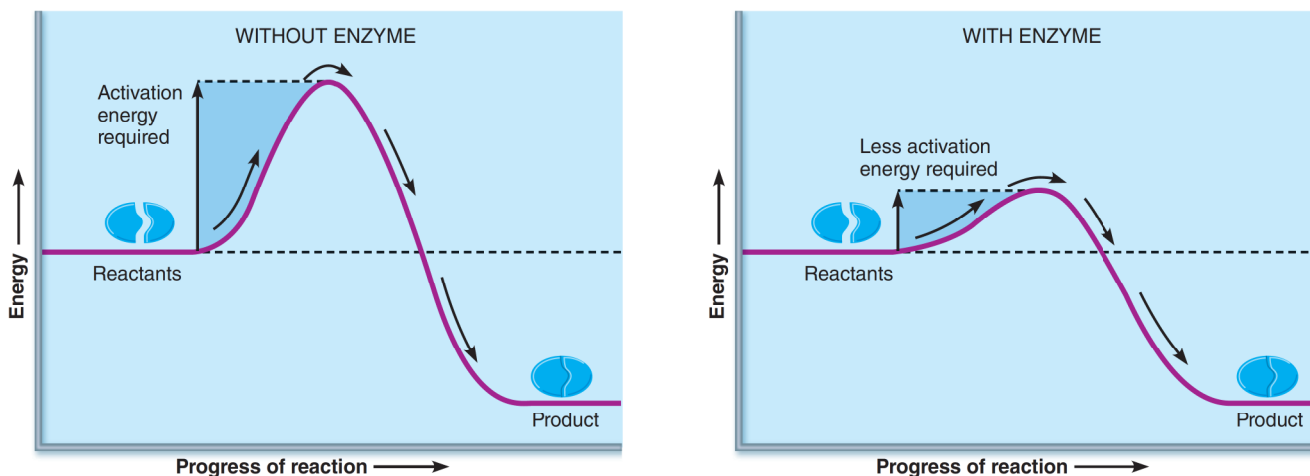


Figure 2.20 Enzymes lower the activation energy required for a reaction.

reactions. More specifically, enzymes can be thought of as chemical traffic cops that keep our metabolic pathways flowing. Enzymes cannot force chemical reactions to occur between molecules that would not otherwise react. They can only increase the speed of reaction, and they do so by staggering amounts—from 100,000 to over 1 billion times the rate of an uncatalyzed reaction. Without enzymes, biochemical reactions proceed so slowly that for practical purposes they do not occur at all.

Characteristics of Enzymes Some enzymes are purely protein. In other cases, the functional enzyme consists of two parts, collectively called a **holoenzyme**: an **apoenzyme** (the protein portion) and a **cofactor**. Depending on the enzyme, the cofactor may be an ion of a metal element such as copper or iron, or an organic molecule needed to assist the reaction in some particular way. Most organic cofactors are derived from vitamins (especially the B complex vitamins). This type of cofactor is more precisely called a **coenzyme**.

Each enzyme is chemically specific. Some enzymes control only a single chemical reaction. Others exhibit a broader specificity in that they can bind with molecules that differ slightly and thus regulate a small group of related reactions. The substance on which an enzyme acts is called a **substrate**.

The presence of specific enzymes determines not only which reactions will be speeded up, but also which reactions will occur—no enzyme, no reaction. This also means that unwanted or unnecessary chemical reactions do not occur.

Most enzymes are named for the type of reaction they catalyze. *Hydrolases* (hi'druh-lās-es) add water during hydrolysis reactions, *oxidases* (ok'sī-dās-es) oxidize reactants by adding oxygen or removing hydrogen, and so on. You can recognize most enzyme names by the suffix *-ase*.

In many cases, enzymes are part of cellular membranes in a bucket-brigade type of arrangement. The product of one enzyme-catalyzed reaction becomes the substrate of the neighboring enzyme, and so on. Some enzymes are produced in an inactive form and must be activated in some way before they can function, often by a change in the pH of their surroundings. For example,

digestive enzymes produced in the pancreas are activated in the small intestine, where they actually do their work. If they were produced in active form, the pancreas would digest itself.

Sometimes, enzymes are inactivated immediately after they have performed their catalytic function. This is true of enzymes that promote blood clot formation when the wall of a blood vessel is damaged. Once clotting is triggered, those enzymes are inactivated. Otherwise, you would have blood vessels full of solid blood instead of one protective clot. (Eek!)

Enzyme Action How do enzymes perform their catalytic role? Every chemical reaction requires that a certain amount of energy, called **activation energy**, be absorbed to prime the reaction. The activation energy is the amount of energy needed to break the bonds of the reactants so that they can rearrange themselves and become the product. It is present when kinetic energy pushes the reactants to an energy level where their random collisions are forceful enough to ensure interaction. Activation energy is needed regardless of whether the overall reaction is ultimately energy absorbing or energy releasing.

One way to increase kinetic energy is to increase the temperature, but higher temperatures denature proteins. (This is why a high fever can be a serious event.) Enzymes allow reactions to occur at normal body temperature by decreasing the amount of activation energy required (**Figure 2.20**).

Exactly how do enzymes accomplish this remarkable feat? The answer is not fully understood. However, we know that, due to structural and electrostatic factors, they decrease the randomness of reactions by binding to the reacting molecules temporarily and presenting them to each other in the proper position for chemical interaction (bond making or breaking) to occur.

Three basic steps appear to be involved in enzyme action (**Figure 2.21**).

- 1 **Substrate(s) bind to the enzyme's active site, temporarily forming an enzyme-substrate complex.** Substrate binding causes the active site to change shape so that the substrate and the active site fit together precisely, and in an orientation that favors reaction. Whether the shape of the substrate

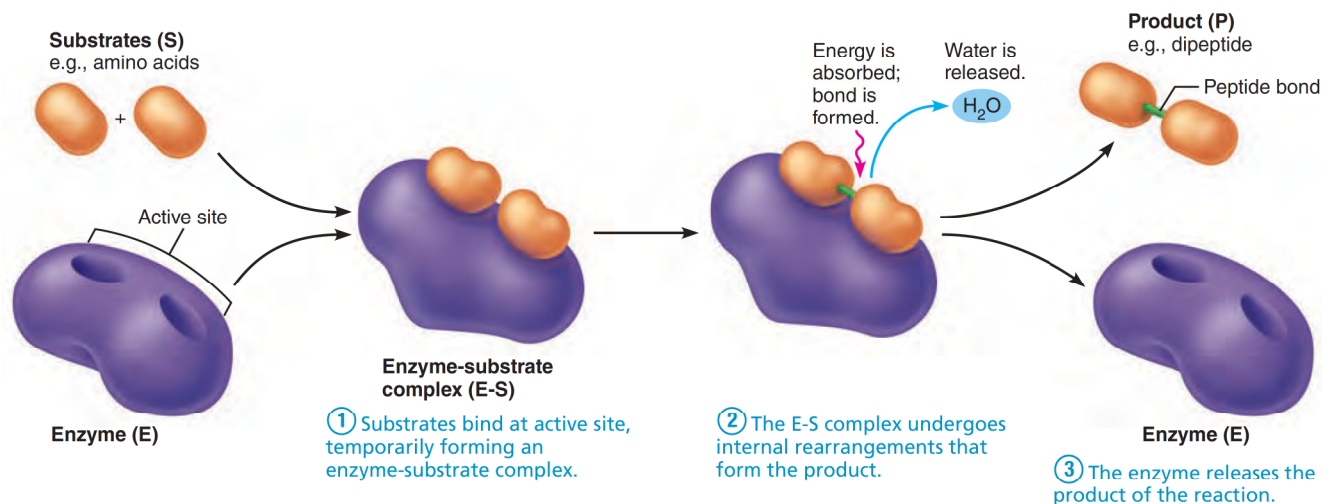


Figure 2.21 Mechanism of enzyme action. In this example, the enzyme catalyzes the formation of a dipeptide from specific amino acids. **Summary:** $E + S \rightarrow E-S \rightarrow P + E$

forces the enzyme to wiggle into a complementary shape or whether the enzyme wiggles around trying several different conformations until one finally fits the substrate is still a question. Although enzymes are specific for particular substrates, other (nonsubstrate) molecules may act as *enzyme inhibitors* if their structure is similar enough to occupy or block the enzyme's active site.

- ② **The enzyme-substrate complex undergoes internal rearrangements that form the product(s).** This step shows the catalytic role of an enzyme.
- ③ **The enzyme releases the product(s) of the reaction.** If the enzyme became part of the product, it would be a reactant and not a catalyst. The enzyme is not changed and returns to its original shape, available to catalyze another reaction.

Because enzymes are unchanged by their catalytic role and can act again and again, cells need only small amounts of each enzyme. Catalysis occurs with incredible speed. Most enzymes can catalyze millions of reactions per minute.

✓ Check Your Understanding

27. What is the main event that molecular chaperones prevent?
28. How do enzymes reduce the amount of activation energy needed to make a chemical reaction go?

For answers, see Appendix H.

Nucleic Acids (DNA and RNA)

- ✓ Compare and contrast DNA and RNA.

The **nucleic acids** (nu-kle'ic), composed of carbon, oxygen, hydrogen, nitrogen, and phosphorus, are the largest molecules in the body. The nucleic acids include two major classes of molecules, **deoxyribonucleic acid (DNA)** (de-ok'si-ri'bo-nu-kle'ik) and **ribonucleic acid (RNA)**.

The structural units of nucleic acids, called **nucleotides**, are quite complex. Each nucleotide consists of three components: a nitrogen-containing base, a pentose sugar, and a phosphate group (**Figure 2.22a**). Five major varieties of nitrogen-containing bases can contribute to nucleotide structure: **adenine**, abbreviated **A** (ad'ě-něn); **guanine**, **G** (gwan'ën); **cytosine**, **C** (si'to-sën); **thymine**, **T** (thi'mën); and **uracil**, **U** (u'rah-sil). Adenine and guanine are large, two-ring bases (called purines), whereas cytosine, thymine, and uracil are smaller, single-ring bases (called pyrimidines).

The stepwise synthesis of a nucleotide involves the attachment of a base to the pentose sugar to form first a *nucleoside*, named for the nitrogenous base it contains. The nucleotide is formed when a phosphate group is bonded to the sugar of the nucleoside.

Although DNA and RNA are both composed of nucleotides, they differ in many respects, as summarized in **Table 2.4** on p. 55. Typically, DNA is found in the nucleus (control center) of the cell, where it constitutes the *genetic material*, also called the *genes*, or more recently the *genome*. DNA has two fundamental roles: It replicates (reproduces) itself before a cell divides, ensuring that the genetic information in the descendant cells is identical, and it provides the basic instructions for building every protein in the body. Although we have said that enzymes govern all chemical reactions, remember that enzymes, too, are proteins formed at the direction of DNA.

By providing the information for protein synthesis, DNA determines what type of organism you will be—frog, human, oak tree—and directs your growth and development. It also accounts for your uniqueness. A technique called DNA fingerprinting can help solve forensic mysteries (for example, verify one's presence at a crime scene), identify badly burned or mangled bodies at a disaster scene, and establish or disprove paternity. DNA fingerprinting analyzes tiny samples of DNA taken from blood, semen, or other body tissues and shows the results as a "genetic barcode" that distinguishes each of us from all others.

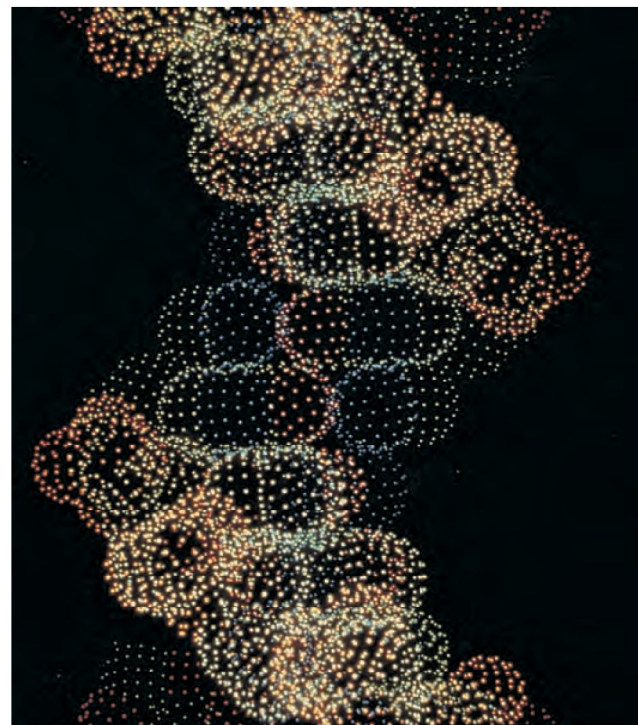
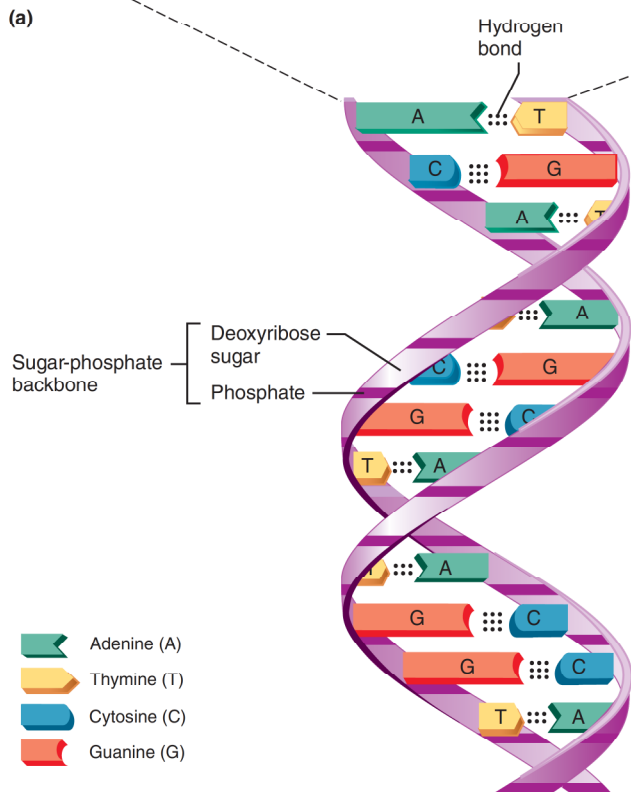
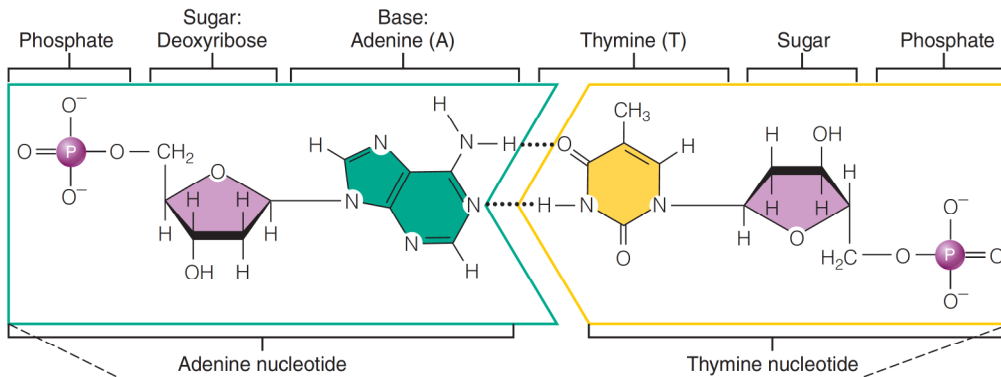


Figure 2.22 Structure of DNA. (a) The unit of DNA is the nucleotide, which is composed of a deoxyribose sugar molecule linked to a phosphate group, with a base attached to the sugar. Two nucleotides, linked by hydrogen bonds between their complementary bases, are illustrated. (b) DNA is a coiled double polymer of nucleotides (a double helix). The backbones of the ladderlike molecule are formed by alternating sugar and phosphate units. The rungs are formed by the binding together of complementary bases (A-T and G-C) by hydrogen bonds (shown by dotted lines). (c) Computer-generated image of a DNA molecule.

DNA is a long, double-stranded polymer—a double chain of nucleotides (Figure 2.22b and c). The bases in DNA are A, G, C, and T, and its pentose sugar is *deoxyribose* (as reflected in its name). Its two nucleotide chains are held together by hydrogen bonds between the bases, so that a ladderlike molecule is formed. Alternating sugar and phosphate components of each chain form the *backbones* or “uprights” of the “ladder,” and the joined bases form the “rungs.” The whole

molecule is coiled into a spiral staircase-like structure called a **double helix**.

Bonding of the bases is very specific: A always bonds to T, and G always bonds to C. A and T are therefore called **complementary bases**, as are C and G. According to these base-pairing rules, ATGA on one DNA nucleotide strand would necessarily be bonded to TACT (a complementary base sequence) on the other strand.

Table 2.4 Comparison of DNA and RNA

CHARACTERISTIC	DNA	RNA
Major cellular site	Nucleus	Cytoplasm (cell area outside the nucleus)
Major functions	Is the genetic material; directs protein synthesis; replicates itself before cell division	Carries out the genetic instructions for protein synthesis
Sugar	Deoxyribose	Ribose
Bases	Adenine, guanine, cytosine, thymine	Adenine, guanine, cytosine, uracil
Structure	Double strand coiled into a double helix	Single strand, straight or folded

RNA is located chiefly outside the nucleus and can be considered a “molecular slave” of DNA. That is, RNA carries out the orders for protein synthesis issued by DNA. [Viruses in which RNA (rather than DNA) is the genetic material are an exception to this generalization.]

RNA molecules are single strands of nucleotides. RNA bases include A, G, C, and U (U replaces the T found in DNA), and its sugar is *ribose* instead of deoxyribose. The three major varieties of RNA (messenger RNA, ribosomal RNA, and transfer RNA) are distinguished by their relative size and shape, and each has a specific role to play in carrying out DNA’s instructions for protein synthesis. In addition to these three RNAs, small RNA molecules called *microRNAs* (*miRNAs*) appear to control genetic expression by shutting down genes or altering their expression. We discuss DNA replication and the relative roles of DNA and RNA in protein synthesis in Chapter 3.

✓ Check Your Understanding

- How do DNA and RNA differ in the bases and sugars they contain?
- What are two important roles of DNA?

For answers, see Appendix H.

Adenosine Triphosphate (ATP)

✓ Explain the role of ATP in cell metabolism.

Glucose is the most important cellular fuel, but none of the chemical energy contained in its bonds is used directly to power cellular work. Instead, energy released during glucose catabolism is coupled to the synthesis of **adenosine triphosphate (ATP)**. In other words, some of this energy is captured and stored as small packets of energy in the bonds of ATP. ATP is the primary energy-transferring molecule in cells and it provides a form of energy that is immediately usable by all body cells.

Structurally, ATP is an adenine-containing RNA nucleotide to which two additional phosphate groups have been added (**Figure 2.23**). Chemically, the triphosphate tail of ATP can be compared to a tightly coiled spring ready to uncoil with tremendous energy when the catch is released. Actually, ATP is a very unstable energy-storing molecule because its three negatively charged phosphate groups are closely packed and repel

each other. When its terminal high-energy phosphate bonds are broken (hydrolyzed), the chemical “spring” relaxes and the molecule as a whole becomes more stable.

Cells tap ATP’s bond energy during coupled reactions by using enzymes to transfer the terminal phosphate groups from ATP to other compounds. These newly *phosphorylated* molecules are said to be “primed” and temporarily become more energetic and capable of performing some type of cellular work. In the process of doing their work, they lose the phosphate group. The amount of energy released and transferred during ATP hydrolysis corresponds closely to that needed to drive most biochemical reactions. As a result, cells are protected from excessive energy release that might be damaging, and energy squandering is kept to a minimum.

Cleaving the terminal phosphate bond of ATP yields a molecule with two phosphate groups—*adenosine diphosphate*

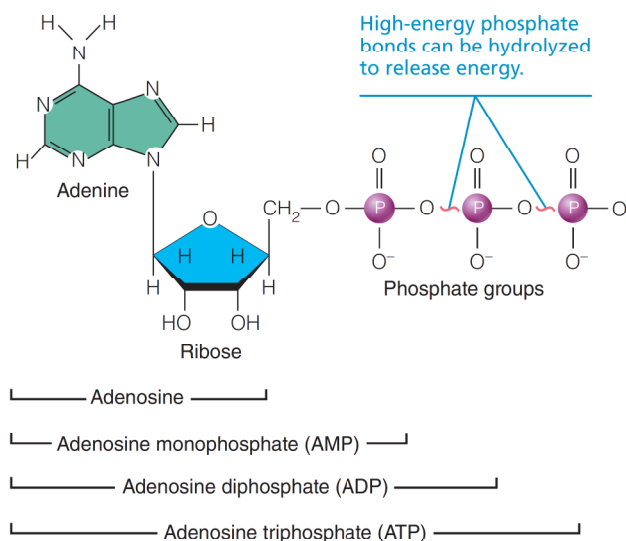


Figure 2.23 Structure of ATP (adenosine triphosphate). ATP is an adenine nucleotide to which two additional phosphate groups have been attached during breakdown of food fuels. When the terminal phosphate group is cleaved off, energy is released to do useful work and ADP (adenosine diphosphate) is formed. When the terminal phosphate group is cleaved off ADP, a similar amount of energy is released and AMP (adenosine monophosphate) is formed.

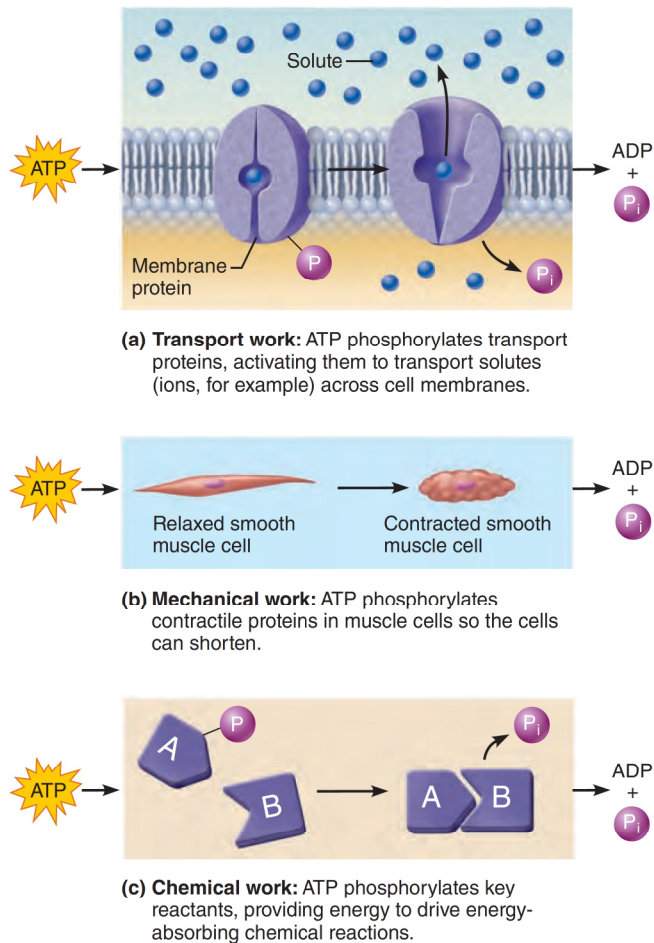


Figure 2.24 Three examples of cellular work driven by energy from ATP.

(ADP)—and an inorganic phosphate group, indicated by P_i , accompanied by a transfer of energy:



As ATP is hydrolyzed to provide energy for cellular needs, ADP accumulates. Cleavage of the terminal phosphate bond of ADP liberates a similar amount of energy and produces adenosine monophosphate (AMP).

The cell's ATP supplies are replenished as glucose and other fuel molecules are oxidized and their bond energy is released. The same amount of energy that is liberated when ATP's terminal phosphates are cleaved off must be captured and used to reverse the reaction to reattach phosphates and re-form the energy-transferring phosphate bonds. Without ATP, molecules cannot be made or degraded, cells cannot transport substances across their membrane boundaries, muscles cannot shorten to tug on other structures, and life processes cease (**Figure 2.24**).

✓ Check Your Understanding

- Glucose is an energy-rich molecule. So why do body cells need ATP?
- What change occurs in ATP when it releases energy?

For answers, see Appendix H.

Chapter Summary

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PART 1 Basic Chemistry

Definition of Concepts: Matter and Energy (pp. 23–25)

Matter (p. 24)

- Matter is anything that takes up space and has mass.

Energy (pp. 24–25)

- Energy is the capacity to do work or put matter into motion.
- Energy exists as potential energy (stored energy or energy of position) and kinetic energy (active or working energy).
- Forms of energy involved in body functioning are chemical, electrical, radiant, and mechanical. Of these, chemical (bond) energy is most important.
- Energy may be converted from one form to another, but some energy is always unusable (lost as heat) in such transformations.

Composition of Matter: Atoms and Elements

(pp. 25–28)

- Elements are unique substances that cannot be decomposed into simpler substances by ordinary chemical methods. Four elements (carbon, hydrogen, oxygen, and nitrogen) make up 96% of body weight.

Atomic Structure (pp. 25–26)

- The building blocks of elements are atoms.
- Atoms are composed of positively charged protons, negatively charged electrons, and uncharged neutrons. Protons and neutrons are located in the atomic nucleus, constituting essentially the atom's total mass. Electrons are outside the nucleus in the electron shells. In any atom, the number of electrons equals the number of protons.

Identifying Elements (pp. 27–28)

- Atoms may be identified by their atomic number (p^+) and mass number ($p^+ + n^0$). The notation ${}^4_2\text{He}$ means that helium (He) has an atomic number of 2 and a mass number of 4.
- Isotopes of an element differ in the number of neutrons they contain. The atomic weight of any element is approximately equal to the mass number of its most abundant isotope.

Radioisotopes (p. 28)

- Many heavy isotopes are unstable (radioactive). These so-called radioisotopes decompose to more stable forms by emitting alpha or beta particles or gamma rays. Radioisotopes are useful in medical diagnosis and treatment and in biochemical research.

How Matter Is Combined: Molecules and Mixtures (pp. 28–30)**Molecules and Compounds** (pp. 28–29)

- A molecule is the smallest unit resulting from the chemical bonding of two or more atoms. If the atoms are different, they form a molecule of a compound.

Mixtures (pp. 29–30)

- Mixtures are physical combinations of solutes in a solvent. Mixture components retain their individual properties.
- The types of mixtures, in order of increasing solute size, are solutions, colloids, and suspensions.
- Solution concentrations are typically designated in terms of percent or molarity.

Distinguishing Mixtures from Compounds (p. 30)

- Compounds are homogeneous; their elements are chemically bonded. Mixtures may be homogeneous or heterogeneous; their components are physically combined and separable.

Chemical Bonds (pp. 30–35)**The Role of Electrons in Chemical Bonding** (p. 31)

- Electrons of an atom occupy areas of space called electron shells or energy levels. Electrons in the shell farthest from the nucleus (valence shell) are most energetic.
- Chemical bonds are energy relationships between valence shell electrons of the reacting atoms. Atoms with a full valence shell or eight valence shell electrons are chemically unreactive (inert). Those with an incomplete valence shell interact with other atoms to achieve stability.

Types of Chemical Bonds (pp. 31–35)

- Ionic bonds are formed when valence shell electrons are completely transferred from one atom to another.
- Covalent bonds are formed when atoms share electron pairs. If the electron pairs are shared equally, the molecule is nonpolar. If they are shared unequally, it is polar (a dipole).

- Hydrogen bonds are weak bonds formed between one hydrogen atom, already covalently linked to an electronegative atom, and another electronegative atom (such as nitrogen or oxygen). They bind together different molecules (e.g., water molecules) or different parts of the same molecule (as in protein molecules).

Chemical Reactions (pp. 35–38)**Chemical Equations** (pp. 35–36)

- Chemical reactions involve the formation, breaking, or rearrangement of chemical bonds.

Patterns of Chemical Reactions (pp. 36–37)

- Chemical reactions are either anabolic (constructive) or catabolic (destructive). They include synthesis, decomposition, and exchange reactions. Oxidation-reduction reactions may be considered a special type of exchange (or decomposition) reaction.

Energy Flow in Chemical Reactions (p. 37)

- Bonds are energy relationships and there is a net loss or gain of energy in every chemical reaction.
- In exergonic reactions, energy is liberated. In endergonic reactions, energy is absorbed.

Reversibility of Chemical Reactions (p. 37)

- If reaction conditions remain unchanged, all chemical reactions eventually reach a state of chemical equilibrium in which the reaction proceeds in both directions at the same rate.
- All chemical reactions are theoretically reversible, but many biological reactions go in only one direction because of energy requirements or the removal of reaction products.

Factors Influencing the Rate of Chemical Reactions (pp. 37–38)

- Chemical reactions occur only when particles collide and valence shell electrons interact.
- The smaller the reacting particles, the greater their kinetic energy and the faster the reaction rate. Higher temperature or reactant concentration, as well as the presence of catalysts, increases chemical reaction rates.

PART 2 Biochemistry**Inorganic Compounds** (pp. 38–41)

- Most inorganic compounds do not contain carbon. Those found in the body include water, salts, and inorganic acids and bases.

Water (pp. 38–39)

- Water is the single most abundant compound in the body. It absorbs and releases heat slowly, acts as a universal solvent, participates in chemical reactions, and cushions body organs.

Salts (p. 39)

- Salts are ionic compounds that dissolve in water and act as electrolytes. Calcium and phosphorus salts contribute to the hardness of bones and teeth. Ions of salts are involved in many physiological processes.

Acids and Bases (pp. 39–41)

- Acids are proton donors; in water, they ionize and dissociate, releasing hydrogen ions (which account for their properties) and anions.

iP Fluid, Electrolyte and Acid/Base Balance; Topic: Acid Base Homeostasis, pp. 1–12, 16, 17.

5. Bases are proton acceptors. The most common inorganic bases are the hydroxides; bicarbonate ion and ammonia are important bases in the body.
6. pH is a measure of hydrogen ion concentration of a solution (in moles per liter). A pH of 7 is neutral; a higher pH is alkaline, and a lower pH is acidic. Normal blood pH is 7.35–7.45. Buffers help to prevent excessive changes in the pH of body fluids.

iP Fluid, Electrolyte and Acid/Base Balance; Topic: Introduction to Body Fluids, pp. 1–8.**Organic Compounds (pp. 41–56)**

1. Organic compounds contain carbon. Those found in the body include carbohydrates, lipids, proteins, and nucleic acids, all of which are synthesized by dehydration synthesis and digested by hydrolysis. All of these biological molecules contain C, H, and O. Proteins and nucleic acids also contain N.

Carbohydrates (p. 43)

2. Carbohydrate building blocks are monosaccharides, the most important of which are hexoses (glucose, fructose, galactose) and pentoses (ribose, deoxyribose).
3. Disaccharides (sucrose, lactose, maltose) and polysaccharides (starch, glycogen) are composed of linked monosaccharide units.
4. Carbohydrates, particularly glucose, are the major energy fuel for forming ATP. Excess carbohydrates are stored as glycogen or converted to fat for storage.

Lipids (pp. 43–47)

5. Lipids dissolve in fats or organic solvents, but not in water.
6. Triglycerides are composed of fatty acid chains and glycerol. They are found chiefly in fatty tissue where they provide insulation and reserve body fuel. Unsaturated fatty acid chains produce oils. Saturated fatty acids produce solid fats typical of animal fats.
7. Phospholipids are modified phosphorus-containing triglycerides that have polar and nonpolar portions. They are found in all plasma membranes.
8. The steroid cholesterol is found in cell membranes and is the basis of steroid hormones, bile salts, and vitamin D.

Proteins (pp. 47–53)

9. The unit of proteins is the amino acid, and 20 common amino acids are found in the body.

10. Many amino acids joined by peptide bonds form a polypeptide. A protein (one or more polypeptides) is distinguished by the number and sequence of amino acids in its chain(s) and by the complexity of its three-dimensional structure.
11. Fibrous proteins, such as keratin and collagen, have secondary (α -helix or β -pleated sheet) and perhaps tertiary and quaternary structure. Fibrous proteins are used as structural materials.
12. Globular proteins achieve tertiary and sometimes quaternary structure and are generally spherical, soluble molecules. Globular proteins (e.g., enzymes, some hormones, antibodies, hemoglobin) perform special functional roles for the cell (e.g., catalysis, molecule transport).
13. Proteins are denatured by extremes of temperature or pH. Denatured globular proteins are unable to perform their usual function.
14. Molecular chaperones assist in folding proteins into their functional 3-D shape. They are synthesized in greater amounts when cells are stressed by environmental factors.
15. Enzymes are biological catalysts. They increase the rate of chemical reactions by decreasing the amount of activation energy needed. They do this by combining with the reactants and holding them in the proper position to interact. Many enzymes require cofactors to function.

Nucleic Acids (DNA and RNA) (pp. 53–55)

16. Nucleic acids include deoxyribonucleic acid (DNA) and ribonucleic acid (RNA). The structural unit of nucleic acids is the nucleotide, which consists of a nitrogenous base (adenine, guanine, cytosine, thymine, or uracil), a sugar (ribose or deoxyribose), and a phosphate group.
17. DNA is a double-stranded helix. It contains deoxyribose and the bases A, G, C, and T. DNA specifies protein structure and replicates itself exactly before cell division.
18. RNA is single stranded. It contains ribose and the bases A, G, C, and U. RNAs involved in carrying out DNA's instructions for protein synthesis include messenger, ribosomal, and transfer RNA.

Adenosine Triphosphate (ATP) (pp. 55–56)

19. ATP is the universal energy compound of body cells. Some of the energy liberated by the breakdown of glucose and other food fuels is captured in the bonds of ATP molecules and transferred via coupled reactions to energy-consuming reactions.

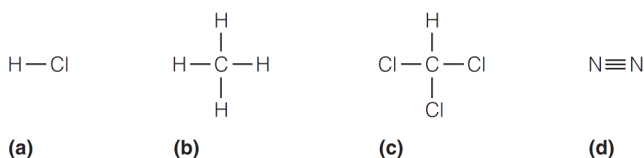
Review Questions

Multiple Choice/Matching

(Some questions have more than one correct answer. Select the best answer or answers from the choices given.)

1. Which of the following forms of energy is the stimulus for vision? (a) chemical, (b) electrical, (c) mechanical, (d) radiant.
2. All of the following are examples of the four major elements contributing to body mass except (a) hydrogen, (b) carbon, (c) nitrogen, (d) sodium, (e) oxygen.
3. The mass number of an atom is (a) equal to the number of protons it contains, (b) the sum of its protons and neutrons, (c) the sum of all of its subatomic particles, (d) the average of the mass numbers of all of its isotopes.
4. A deficiency in this element can be expected to reduce the hemoglobin content of blood: (a) Fe, (b) I, (c) F, (d) Ca, (e) K.
5. Which set of terms best describes a proton? (a) negative charge, massless, in the orbital; (b) positive charge, 1 amu, in the nucleus; (c) uncharged, 1 amu, in the nucleus.
6. The subatomic particles responsible for the chemical behavior of atoms are (a) electrons, (b) ions, (c) neutrons, (d) protons.
7. In the body, carbohydrates are stored in the form of (a) glycogen, (b) starch, (c) cholesterol, (d) polypeptides.

- Which of the following does *not* describe a mixture?
 - properties of its components are retained,
 - chemical bonds are formed,
 - components can be separated physically,
 - includes both heterogeneous and homogeneous examples.
- In a beaker of water, the water-water bonds can properly be called
 - ionic bonds,
 - polar covalent bonds,
 - nonpolar covalent bonds,
 - hydrogen bonds.
- When a pair of electrons is shared between two atoms, the bond formed is called
 - a single covalent bond,
 - a double covalent bond,
 - a triple covalent bond,
 - an ionic bond.
- Molecules formed when electrons are shared unequally are
 - salts,
 - polar molecules,
 - nonpolar molecules.
- Which of the following covalently bonded molecules are polar?



- Identify each reaction as one of the following: (a) a synthesis reaction, (b) a decomposition reaction, (c) an exchange reaction.
 - $2\text{Hg} + \text{O}_2 \rightarrow 2\text{HgO}$
 - $\text{HCl} + \text{NaOH} \rightarrow \text{NaCl} + \text{H}_2\text{O}$
- Factors that accelerate the rate of chemical reactions include all but
 - the presence of catalysts,
 - increasing the temperature,
 - increasing the particle size,
 - increasing the concentration of the reactants.
- Which of the following molecules is an inorganic molecule?
 - sucrose,
 - cholesterol,
 - collagen,
 - sodium chloride.
- Water's importance to living systems reflects
 - its polarity and solvent properties,
 - its high heat capacity,
 - its high heat of vaporization,
 - its chemical reactivity,
 - all of these.
- Acids
 - release hydroxyl ions when dissolved in water,
 - are proton acceptors,
 - cause the pH of a solution to rise,
 - release protons when dissolved in water.
- A chemist, during the course of an analysis, runs across a chemical composed of carbon, hydrogen, and oxygen in the proportion 1:2:1 and having a six-sided molecular shape. It is probably
 - a pentose,
 - an amino acid,
 - a fatty acid,
 - a monosaccharide,
 - a nucleic acid.
- A triglyceride consists of
 - glycerol plus three fatty acids,
 - a sugar-phosphate backbone to which two amino groups are attached,
 - two to several hexoses,
 - amino acids that have been thoroughly saturated with hydrogen.
- A chemical has an amine group and an organic acid group. It does not, however, have any peptide bonds. It is
 - a monosaccharide,
 - an amino acid,
 - a protein,
 - a fat.
- The lipid(s) used as the basis of vitamin D, sex hormones, and bile salts is/are
 - triglycerides,
 - cholesterol,
 - phospholipids,
 - prostaglandin.
- Enzymes are organic catalysts that
 - alter the direction in which a chemical reaction proceeds,
 - determine the nature of the products of a reaction,
 - increase the speed of a chemical reaction,
 - are essential raw materials for a chemical reaction that are converted into some of its products.

Short Answer Essay Questions

- Define or describe energy, and explain the relationship between potential and kinetic energy.

- Some energy is lost in every energy conversion. Explain the meaning of this statement. (Direct your response to answering the question: Is it really lost? If not, what then?)
- Provide the atomic symbol for each of the following elements:
 - calcium,
 - carbon,
 - hydrogen,
 - iron,
 - nitrogen,
 - oxygen,
 - potassium,
 - sodium.
- Consider the following information about three atoms:

$${}^{12}_6\text{C} \quad {}^{13}_6\text{C} \quad {}^{14}_6\text{C}$$
 - How are they similar to one another? (b) How do they differ from one another? (c) What are the members of such a group of atoms called? (d) Using the planetary model, draw the atomic configuration of ${}^{12}_6\text{C}$ showing the relative position and numbers of its subatomic particles.
- How many moles of aspirin, $\text{C}_9\text{H}_8\text{O}_4$, are in a bottle containing 450 g by weight? (*Note:* The approximate atomic weights of its atoms are C = 12, H = 1, and O = 16.)
- Given the following types of atoms, decide which type of bonding, ionic or covalent, is most likely to occur: (a) two oxygen atoms; (b) four hydrogen atoms and one carbon atom; (c) a potassium atom (${}^{39}_{19}\text{K}$) and a fluorine atom (${}^{19}_9\text{F}$).
- What are hydrogen bonds and how are they important in the body?
- The following equation, which represents the oxidative breakdown of glucose by body cells, is a reversible reaction.

$$\text{Glucose} + \text{oxygen} \rightarrow \text{carbon dioxide} + \text{water} + \text{ATP}$$
 - How can you indicate that the reaction is reversible? (b) How can you indicate that the reaction is in chemical equilibrium? (c) Define chemical equilibrium.
- Differentiate clearly between primary, secondary, and tertiary protein structure.
- Dehydration and hydrolysis reactions are essentially opposite reactions. How are they related to the synthesis and degradation (breakdown) of biological molecules?
- Describe the mechanism of enzyme action.
- Explain the importance of molecular chaperones.
- Explain why, if you pour water into a glass very carefully, you can "stack" the water slightly above the rim of the glass.



Critical Thinking and Clinical Application Questions

- As Ben jumped on his bike and headed for the freshwater lake, his mother called after him, "Don't swim if we have an electrical storm—it looks threatening." This was a valid request. Why?
- Some antibiotics act by binding to certain essential enzymes in the target bacteria. (a) How might these antibiotics influence the chemical reactions controlled by the enzymes? (b) What is the anticipated effect on the bacteria? On the person taking the antibiotic prescription?
- Mrs. Roberts, in a diabetic coma, has just been admitted to Noble Hospital. Her blood pH indicates that she is in severe acidosis, and measures are quickly instituted to bring her blood pH back within normal limits. (a) Define pH and note the normal pH of blood. (b) Why is severe acidosis a problem?
- Jimmy, a 12-year-old boy, was awakened suddenly by a loud crash. As he sat up in bed, straining to listen, his fright was revealed by his rapid breathing (hyperventilation), a breathing pattern effective in ridding the blood of CO_2 . At this point, was his blood pH rising or falling?
- After you eat a protein bar, which chemical reactions introduced in this chapter must occur for the amino acids in the protein bar to be converted into proteins in your body cells?

AT THE CLINIC

Related Clinical Terms

Acidosis (as^uˈi-doʻsis; *acid* = sour, sharp) A condition of acidity or low pH (below 7.35) of the blood; high hydrogen ion concentration.

Alkalosis (al^uˈkah-loʻsis) A condition of basicity or high pH (above 7.45) of the blood; low hydrogen ion concentration.

Heavy metals Metals with toxic effects on the body, including arsenic, mercury, and lead. Iron, also included in this group, is toxic in high concentrations.

Ionizing radiation Radiation that causes atoms to ionize; for example, radioisotope emissions and X rays.

Ketosis (ke-toʻsis) A condition resulting from excessive ketones (breakdown products of fats) in the blood; common during starvation and acute attacks of diabetes mellitus.

Radiation sickness Disease resulting from exposure of the body to radioactivity; digestive system organs are most affected.