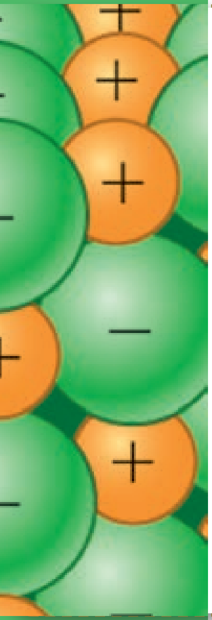
2

Chemistry in Context: The Molecules of Life



Major Themes

- All matter is composed of atoms.
- Atoms are bound together by chemical bonds to form molecules.
- Chemical bonds store energy.
- The electrical charge of atoms and molecules is important in some chemical reactions.
- Chemical reactions create or break chemical bonds.
- Water is an important chemical in some chemical reactions.
- The body strictly controls the acidity of body tissues and fluids.
- Organic molecules are large, complex, and made only by living things.

Chapter Objectives

The Elements of Life 32

1. Give some examples of trace elements and bulk elements.

The Form and Function of Atoms 33

2. Use the periodic table to determine the atomic number and mass of an atom and then calculate the number of protons, electrons, and neutrons.
3. For a given atom, predict the effect of changing the number of protons, electrons, and neutrons.

Chemical Bonds 40

4. Use the electron shell model to explain the formation of ions and ionic bonds, nonpolar covalent bonds, and polar covalent bonds.
5. Explain why only polar molecules and ions can dissolve in water.
6. Explain the importance of functional groups, and list five examples.
7. Explain why an increased concentration of hydrogen or bicarbonate ions results in increased carbon dioxide production, using chemical symbols to illustrate the chemical reaction.

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8. Explain the difference between organic and inorganic molecules.
9. List five important properties of water.
10. Relate the pH scale to changes in hydrogen ion concentrations.
11. Illustrate how different macromolecules (carbohydrates, lipids, proteins, and nucleotides) are assembled from different subunits.
12. Explain how water is involved in dehydration synthesis and hydrolysis.
13. Explain why lipids are hydrophobic but carbohydrates are hydrophilic.
14. Describe the functions of nucleic acid and adenosine triphosphate (ATP).

Chemistry in Context: The Case of Joe G. 51

15. Use the case study to illustrate the importance of buffers.

Case Study: “They made me drink pure lemon juice.”

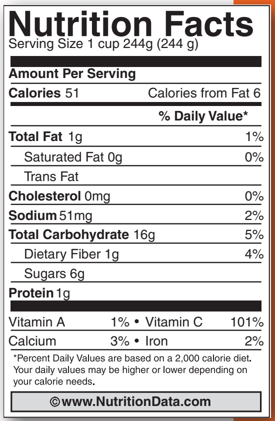
As you read through the following case study, assemble a list of the terms and concepts you must learn in order to understand Joe’s case.

Clinical history: Joe G. is an 18-year-old college freshman in otherwise good health who was deposited in the emergency room by friends who left without providing any information. On questioning by an emergency physician, Joe said, “They made me drink pure lemon juice” and described a fraternity hazing ritual in which he drank most of the contents of three large bottles of commercial lemon juice. He said he vomited once, “but not much came up.”

Physical examination and other data: He was breathing rapidly (24 respirations per minute) and seemed drowsy. He complained of nausea and shortness of breath, but no other physical abnormalities were present. A blood specimen was obtained and analyzed for acidity (pH), bicarbonate, oxygen, and carbon dioxide.

● pH	7.26	Normal, 7.40
● Bicarbonate (HCO_3^-)	21	Normal, 24
● Oxygen	96	Normal, over 80
● Carbon dioxide (CO_2)	33	Normal, 40

Clinical course: A diagnosis of acidosis was made and Joe was given an intravenous solution containing bicarbonate and electrolytes, which was continued overnight. The next morning his respirations were 14 per minute, his laboratory tests were normal, and he was discharged.



Nutrition Facts
Serving Size 1 cup 244g (244 g)

Amount Per Serving	
Calories 51	Calories from Fat 6
	% Daily Value*
Total Fat 1g	1%
Saturated Fat 0g	0%
Trans Fat	
Cholesterol 0mg	0%
Sodium 51mg	2%
Total Carbohydrate 16g	5%
Dietary Fiber 1g	4%
Sugars 6g	
Protein 1g	
Vitamin A 1% • Vitamin C	101%
Calcium 3% • Iron	2%

*Percent Daily Values are based on a 2,000 calorie diet. Your daily values may be higher or lower depending on your calorie needs.

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Nutrition facts panel from a bottle of lemon juice.

In ⬅ Chapter 1, we thought of ourselves as "viewing the landscape of anatomy and physiology from the window of a jet plane." In this chapter, we can think of ourselves on hands and knees, peering through a magnifying glass, studying "the smallest things" our eyes can see: the rocks, roots, and soil of which the landscape is formed. The view from high above and the view at very close range are necessary for us to understand what we will see in the remainder of our journey, which will be, so to speak, on foot, hiking through the hills and dales of human form and function.

Recall that in ⬅ Chapter 1, we introduced you to the levels of organization in the body. Now we are going to discuss the organization of the smallest things: molecules and the atoms of which they are made.

Nothing exists but atoms and empty space.

Democritus, Greek philosopher, circa 470 to 380 BC, who gave us the word *atom*, which in Greek means "indivisible"; that is, the smallest thing that cannot be divided into something even smaller.

The Elements of Life

What are we made of? Although an old nursery rhyme suggests that girls are made of sugar and spice and boys of puppy dogs' tails, we know that all humans are composed of the same fundamental chemicals in roughly the same proportions (Fig. 2.1). These fundamental chemicals, such as oxygen, hydrogen, and nitrogen, are examples of *elements*. Not only human beings but all matter is composed of elements. An **element** is a substance that cannot be reduced to a simpler substance by normal forces such as heat, electricity, magnetism, or a chemical reaction. For example, the element iron always remains iron in chemical reactions; it does not change into another metal such as copper or lead; and elemental iron is the same substance whether it occurs in human blood or in a cast-iron skillet.

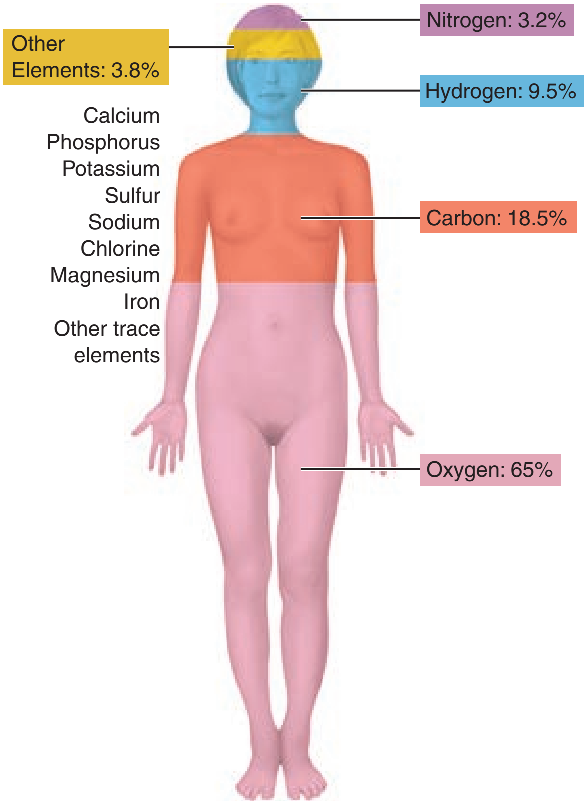
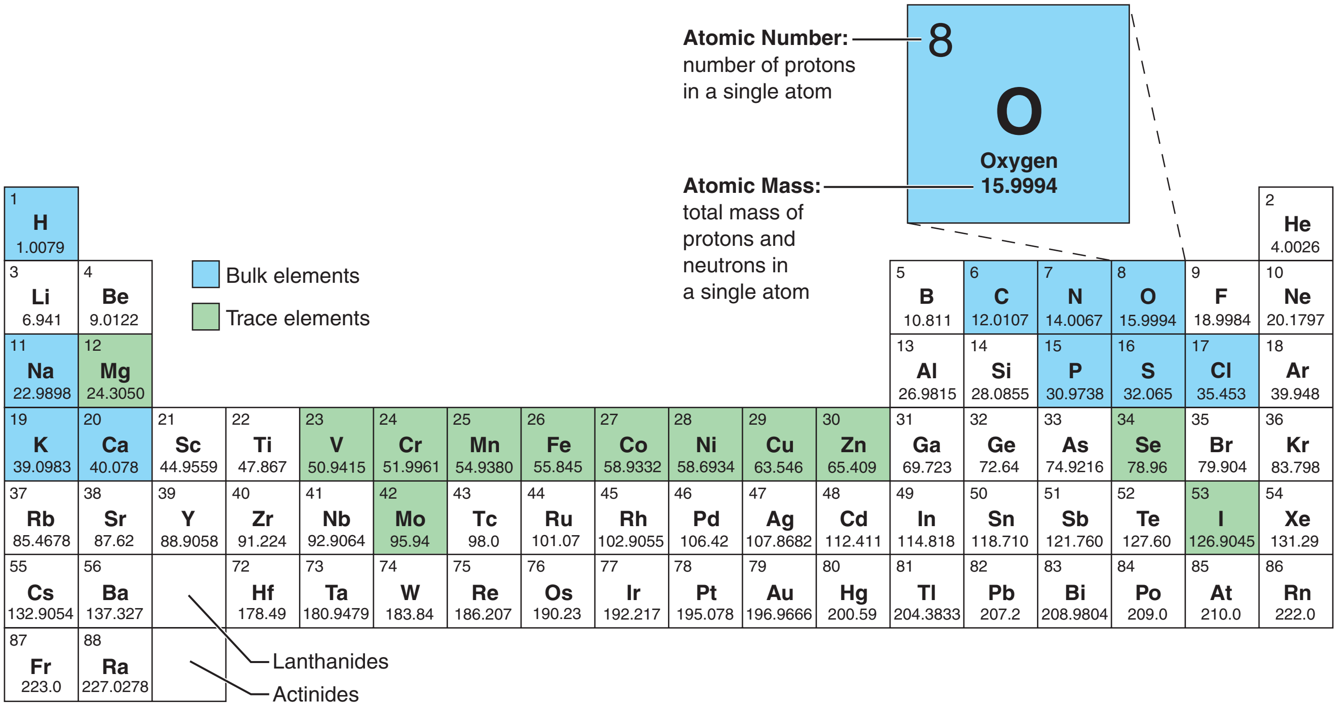


Figure 2.1. Elements of the body. The percentage of body mass made up by each element is shown on the figure. *What is the third most abundant element in the human body?*

Remember This! **Many breakfast cereals are supplemented with iron. You can run an experiment to see whether your favorite cereal contains iron. First crush it up, then stir it with a magnet. If iron is present, bits of cereal should stick to the magnet.**

Over 90 elements are found in nature; about 30 of these are essential for life. Figure 2.2 presents the **periodic table,** a chart illustrating the essential characteristics of all elements known to exist. Each element is listed by its abbreviation (H for hydrogen, for instance). The elements are numbered from the smallest (1, H, hydrogen) to the very large (88, Ra, radium). Elements in the same column tend to share chemical properties. A bit further on, we discuss the logic of the different numbers and the arrangements of the elements into rows.

The four most abundant elements, which constitute 99% of the mass of body tissues, are oxygen (O), carbon (C), hydrogen (H), and nitrogen (N). These elements and a few others are the *bulk elements* (Fig. 2.2), which must be obtained daily in the diet in large quantities. *Trace elements* (*trace minerals, micronutrients*) are chemical elements required for life but only in very small amounts, ordinarily far less than 0.10 g per day. They, too, are obtained through the diet and include cobalt (Co), copper (Cu), chromium (Cr), iron (Fe), iodine (I), selenium (Se), and others. Iron, for instance, is an integral part of hemoglobin, the chemical that carries oxygen in blood.



🟦 Bulk elements

🟩 Trace elements

🟦 1 H 1.0079																	2 He 4.0026
3 Li 6.941	4 Be 9.0122											5 B 10.811	🟦 6 C 12.0107	🟦 7 N 14.0067	🟦 8 O 15.9994	9 F 18.9984	10 Ne 20.1797
🟦 11 Na 22.9898	🟩 12 Mg 24.3050											13 Al 26.9815	14 Si 28.0855	🟦 15 P 30.9738	🟦 16 S 32.065	🟦 17 Cl 35.453	18 Ar 39.948
🟦 19 K 39.0983	🟦 20 Ca 40.078	21 Sc 44.9559	22 Ti 47.867	🟩 23 V 50.9415	🟩 24 Cr 51.9961	🟩 25 Mn 54.9380	🟩 26 Fe 55.845	🟩 27 Co 58.9332	🟩 28 Ni 58.6934	🟩 29 Cu 63.546	🟩 30 Zn 65.409	31 Ga 69.723	32 Ge 72.64	33 As 74.9216	🟩 34 Se 78.96	35 Br 79.904	36 Kr 83.798
37 Rb 85.4678	38 Sr 87.62	39 Y 88.9058	40 Zr 91.224	41 Nb 92.9064	🟩 42 Mo 95.94	43 Tc 98.0	44 Ru 101.07	45 Rh 102.9055	46 Pd 106.42	47 Ag 107.8682	48 Cd 112.411	49 In 114.818	50 Sn 118.710	51 Sb 121.760	52 Te 127.60	🟩 53 I 126.9045	54 Xe 131.29
55 Cs 132.9054	56 Ba 137.327		72 Hf 178.49	73 Ta 180.9479	74 W 183.84	75 Re 186.207	76 Os 190.23	77 Ir 192.217	78 Pt 195.078	79 Au 196.9666	80 Hg 200.59	81 Tl 204.3833	82 Pb 207.2	83 Bi 208.9804	84 Po 209.0	85 At 210.0	86 Rn 222.0
87 Fr 223.0	88 Ra 227.0278																

Figure 2.2. The periodic table. Bulk elements are shown in blue; trace elements are shown in green. *Is chromium (Cr) a bulk element or a trace element?*

Case Note

2.1. As shown on the Nutrition Facts panel in the case study box, lemon juice contains significant amounts of calcium (Ca) and iron (Fe). Which of these is a bulk element and which is a trace element?

People consuming a varied diet will meet their need for most bulk elements because they are found in a wide range of foods. In contrast, trace elements are found only in specific foods, so inadequate intake is more common. When people do not meet their daily need for a particular trace element, a deficiency may develop. Iron deficiency is the most common dietary deficiency in the western world.

To avoid such deficiencies, many basic foods, such as cereals and bread, are supplemented with trace elements such as iron.

2.1 Name the four most abundant elements in the human body.

2.2 True or false. Iron is a trace element.

2.3 True or false. The daily requirement for each trace element is about 1 g.

The Form and Function of Atoms

The average 70-kg (155-lb) human contains 12.6 kg (28 lb) of the element carbon. The carbon is, of course, not gathered into a large chunk—it is divided into approximately $7.0*10^{26}$ (roughly a billion billion billion) carbon *atoms*. An **atom** is the smallest particle of an element that behaves like the element; for example, a carbon atom is the smallest particle of the element carbon that still acts like carbon.

Atoms Are Composed of Subatomic Particles

For over 2,000 years, science has been looking for "the smallest thing" that makes up matter, and in the early 19th century the existence of atoms was proved. However, experiments soon revealed that atoms have an internal structure composed of even smaller *subatomic particles—neutrons, protons, and electrons* (Fig. 2.3A). The behavior of these subatomic particles is key to all physiological functions, from the beating of the heart to the formation of urine. Later discoveries proved that subatomic particles are themselves composed of even tinier *elementary particles* called *quarks*, which now

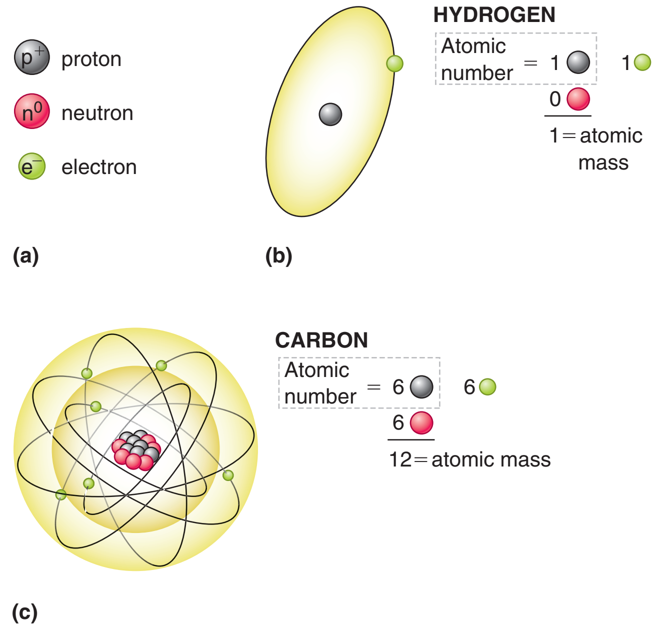


Figure 2.3. Structure of atoms. A. Atoms are composed of protons, neutrons, and electrons. All protons are identical, as are all neutrons, and all electrons. **B.** A hydrogen atom consists of a single electron orbiting a single proton. **C.** The heavier carbon atom contains more protons and electrons, as well as neutrons. Note that the illustrations are NOT to scale—the electrons are much smaller and further away from the nucleus than shown here. *Which atom contains an equal number of protons and neutrons—carbon or hydrogen?*

seem to be the "smallest thing." There are six different types, or *flavors*, of quarks, which combine in different proportions to form subatomic particles.

It is convenient to think of atoms and their subatomic particles as miniature solar systems composed of a central nucleus (the sun), about which other subatomic particles (electrons) orbit like planets (Fig. 2.3). The nucleus of an atom sits in its center and is composed of *protons* and *neutrons*, which are heavy. Surrounding the nucleus are *electrons*, which are very lightweight. Because the nucleus is composed of two heavy particles (protons and neutrons) and electrons are so light, an atom can also be likened to a bowling ball of heavy protons and neutrons with a few gnatlike electrons buzzing around it at a great distance. It is important to realize that this "planetary model" is merely a useful way of thinking about the behavior of atoms because it explains *some* atomic behavior, the behavior that concerns us in this chapter. The *actual* behavior of atoms is very unusual, so unusual that it led Sir Arthur Eddington (1882–1944), an English astronomer, to say that the behavior of atoms and subatomic particles is "…not only stranger than we imagine, it is stranger than we *can* imagine."

Protons and electrons have an *electric charge*; that is, they possess electrical **energy** or the ability to perform work (the accompanying Basic Form, Basic Function box, titled *Energy and Life,* offers more information about the nature of energy). Electric charge is easier to demonstrate than to define—suffice it to say that two opposing charges exist in nature: one called *negative* (−), and its opposite, *positive* (+). By convention, the electron's charge is considered negative, and the proton's charge is positive. The charge of a single electron is designated −1, whereas the charge of a proton is designated +1. The neutron does not possess electrical energy; its charge is 0. If you've ever experimented with magnets, you know that opposite charges attract one another; that is, negative particles are attracted to positive ones and vice versa. In contrast, like charges repel one another: negative particles are repelled by other negative particles; and positives repel positives.

Although the three types of subatomic particles differ in their location, weight, and charge, every particular subatomic particle of the same type is identical: that is, all protons are alike, all electrons are alike, and all neutrons are alike. Therefore, the properties of an atom depend on the number of protons, neutrons, and electrons it contains. For example, hydrogen (H) is a very lightweight gas, and carbon (symbolized by C) is the hard substance that makes up pencil lead. The physical and chemical differences between them are due to their differing numbers of protons, neutrons, and electrons: hydrogen has a nucleus of one proton and no neutrons and is orbited by a single electron (Fig. 2.3B); carbon, on the other hand, is composed of six protons, six neutrons, and six electrons (Fig. 2.3C).

Differences in the number of protons determine the physical and chemical character of the atom and the element it represents. All atoms of a particular element have the same number of protons. The number of protons in the nucleus defines the element: its **atomic number.** Changing the number of protons changes one element into another. For example, by definition an iron atom is one that has a nucleus that contains 26 protons; the number of neutrons and electrons can vary somewhat, but it remains iron only as long as it contains 26 protons. Every atom in iron, therefore, contains 26 protons, and the *atomic number* of iron is 26. Were we to take away one proton, the atoms would become manganese (Mn), a metal defined as having 25 protons (or atomic number 25); or, if we were to add one proton, we would produce cobalt (Co), which by definition is an atom with 27 protons (atomic number 27). Manganese and cobalt have physical and chemical characteristics that are distinctly different from those of iron.

BASIC FORM, BASIC FUNCTION

Energy and Life

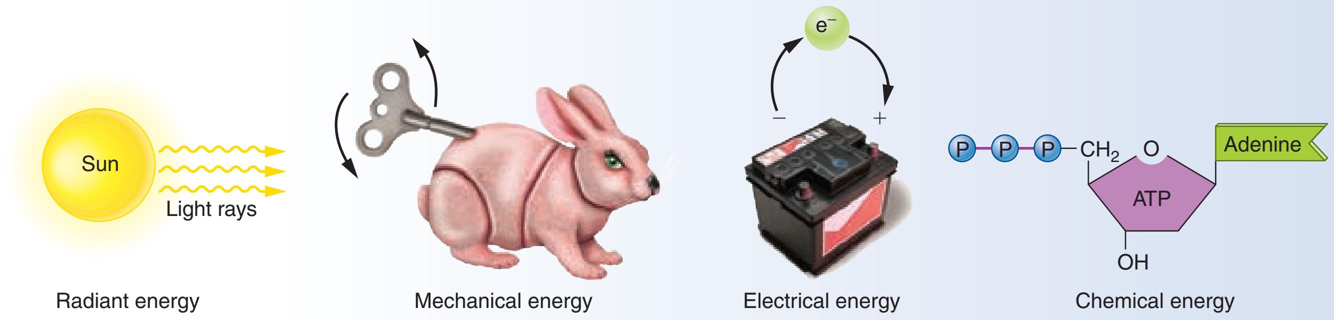
The Greek philosopher Democritus said, "*Nothing exists but atoms and empty space.*" But that's not exactly true—**energy** occupies "empty space" but has no mass (or weight). Although it is not a physical thing like an atom or an axe, it is as real as sunlight. It is *not* nothing.

Energy is observable only by the *effect* it has on matter. For example, light waves are a form of energy that influences every aspect of life on earth. Sunlight heats the earth; without it earth would be a ball of ice. Sugar cane plants capture sunlight and use it to power the synthesis of sugar, which is then consumed by humans and other animals and converted back into energy that powers chemical reactions and maintains body temperature.

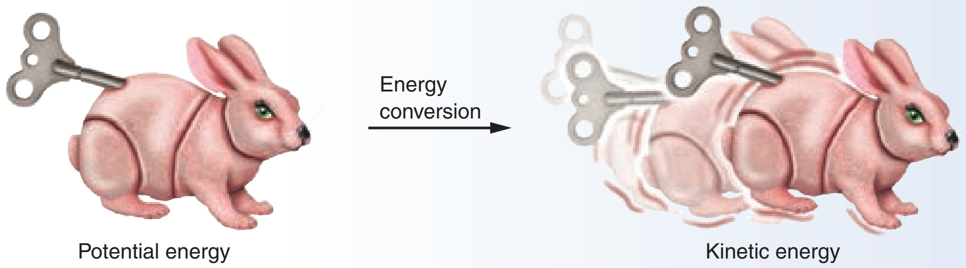
Thus energy and matter are intimately related. Energy can be converted into matter, as in the example of sugar cane sugar, and matter can be converted into energy, as the heat from a fire proves. As the wood is transformed into ashes and smoke, the energy contained in the wood is released as light and heat. After the fire has burned out, the total mass of the wood placed into the fire is equal to the mass of ashes and smoke plus the energy released as heat and light.

Energy can put matter into motion or can stop its motion. It can also be emitted as waves of sound or light that travel through space. Energy exists in several forms:

- *Radiant energy* is emitted as *atomic particles* or as *waves* that travel through space. Ordinary sunlight is a form of radiant energy, for example.
- *Mechanical energy* is stored in mechanical systems. For example, the tightly coiled spring that powers a wind-up toy contains *potential energy*, which is converted to *kinetic energy* (the energy of a moving object) as the toy moves.
- *Electric energy* is the energy associated with the movement of electrons. For example, electric currents and lightning bolts are flows of electrons.
- *Chemical energy* is stored in the bonds that hold molecules together and is released as the bonds are broken and parts of the molecule move apart.



(a) Energy takes many forms.



(b) Energy can be converted between different forms.

Energy. A. Energy can take many forms, including radiant, mechanical, electric, and chemical. **B.** Potential energy can be converted into kinetic energy.

(continued)

Energy and Life *(Continued)*

When energy is stored, it is called **potential energy.** For example, the sun is a ball of hydrogen and helium, each of which is a source of potential energy. It uses some of this energy every day as it burns hydrogen to produce light. As just noted, a wind-up toy contains potential mechanical energy in its coiled spring; a battery contains potential electrical energy in the chemicals and metals it contains; and glucose contains potential chemical energy in the bonds that hold the atoms together. When potential energy is used to put something in motion, the energy is called **kinetic energy.**

Energy of one form can be *converted* into another, and *all energy conversions liberate heat*. For example, a lightbulb converts electrical energy into radiant energy (light waves), and in doing so the bulb becomes hot. The same thing happens in the body: when dietary nutrients, such as sugar and fat, are converted into the mechanical action of running, heat is liberated, which raises body temperature.

Remember This! **No two elements have the same atomic number.**

The total number of protons and neutrons in the nucleus determines the weight of an atom—its **atomic mass.** Electrons are so extremely light that they are not considered. Atoms with a high atomic mass are heavy; those with a low mass are light. In notations of elements, a number to the upper left of the symbol indicates the atomic mass of that element, not its atomic number. For example, normal carbon (C) contains six protons and six neutrons and has an atomic mass of 12, which is indicated as ^{12}C (Fig. 2.3C). Hydrogen (^{1}H) is the smallest and lightest element: it exists as a gas and consists of one proton, no neutrons, and one electron (Fig. 2.3B). Toward the other end of the scale, uranium (^{238}U) is a very heavy metal and each of its atoms consists of 92 protons, 146 neutrons, and 92 electrons. Note that the number of protons does not necessarily equal the number of neutrons.

You can discover the atomic number and mass for each element using a periodic table (Fig. 2.2). Note that the atomic number for each element is given to the upper left of its symbol and the atomic mass below it.

Case Note

2.2. The oxygen concentration in Joe's blood was measured. What is the atomic number and atomic mass of one oxygen atom?

Isotopes Differ in the Number of Neutrons in the Nucleus

Atoms with the same number of protons but different numbers of neutrons are known as **isotopes.** For example, the most common form of carbon contains six protons and six neutrons and has an atomic mass of 12 (^{12}C). If, however, we were to add two neutrons to the nucleus, we would create a *heavy* isotope of carbon (^{14}C) because it contains eight, not six, neutrons. The isotope is said to be heavy because it contains more than the normal number of neutrons. Heavy isotopes are not very stable; they tend to break apart into other, more stable atoms. This process releases *ionizing radiation,* which is particles or energy waves that collide with other atoms and destabilize them. Ionizing radiation can kill cells or damage their DNA in a way that causes cancer.

That isotopes release ionizing particles makes them valuable in medicine and physiology because laboratory instruments can detect the radiation they emit. Typically, the isotope is built into a custom-made molecule that is identical to a natural one except that it emits radiation. After entering the body, the radioactive molecule participates in chemical reactions alongside the natural molecules and emits radiation as it does so, which is detected with a laboratory instrument.

Suppose, for example, we want to study the intestinal absorption of a certain dietary nutrient molecule. All we need to do is to make some of the molecule with an isotope in it and have the patient consume it. If the intestine absorbs it, radioactivity will show up in the blood. For another example of the usefulness of radioactivity, see the following page Clinical Snapshot, titled *Radioactivity: The Good, the Bad, and the Ugly*.

What's more, the radiation produced by isotopes can be used to treat cancer. For example, certain cancers may need a particular molecule to survive. Knowing this, medical scientists can insert an isotopic atom into the molecule and administer it to a patient. As cancer cells use the molecule, they are killed by the radiation.

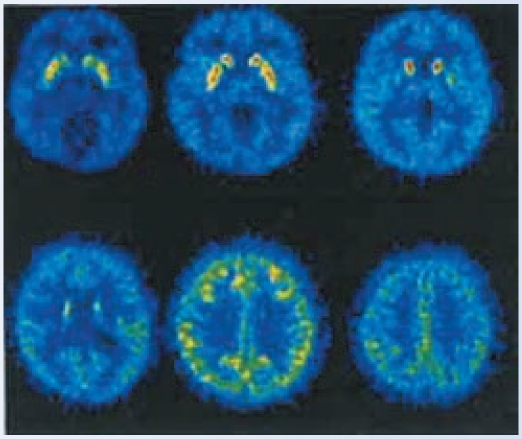
CLINICAL SNAPSHOT

Radioactivity: The Good, the Bad, and the Ugly

"Radioactive." Scary word. Stay away. Handle with extreme caution. Visions of atomic bombs, deadly poisons, and pollution. And all of that is true. For evidence we need look no further than the atomic bombs dropped on Japan in 1945. The blasts in Hiroshima and Nagasaki killed tens of thousands of people immediately and others died of radiation poisoning in the following days, months, and years. Those who survived the shock and heat of the explosion were irradiated by ionizing rays from the blast or were later slowly poisoned by rays from radioactive compounds deposited in the environment. The radiation damaged DNA and caused a great increase in birth defects and cancers in the exposed population that survived. That's "the bad and the ugly" part of radioactivity.

Less well known is the good—in certain well-controlled circumstances radioactive molecules can be used in medical diagnosis and treatment. For example, to make its hormones, the thyroid gland must use tiny amounts of iodine (a trace element). Glands that make too much thyroid hormone (a medical condition called hyperthyroidism) require a lot of iodine. One way to see if the thyroid gland is making too much hormone is to give the patient a tiny dose of radioactive iodine (too little to cause health risk) and monitor the radioactivity that accumulates in the thyroid gland. Patients with hyperthyroidism have very "hot" (radioactive) glands.

Another interesting example is a positron emission tomography (PET) scan using a camera sensitive to radioactive rays instead of rays of light. PET scan technique depends upon the fact that cells burn glucose for energy, and cells that are more active burn more glucose than those that are less active. In the PET procedure, a patient is injected with a tiny dose of radioactive glucose, and images are obtained of various sections of tissue. Physicians can then study which tissues are burning the most glucose—that is, which are metabolically most active. These tissues take up more of the radioactive glucose and, therefore, emit more radioactive waves, which show up as "hot spots" to the PET camera. PET scans can be used to detect cancer because cancer cells multiply rapidly, an activity that requires a lot of glucose.



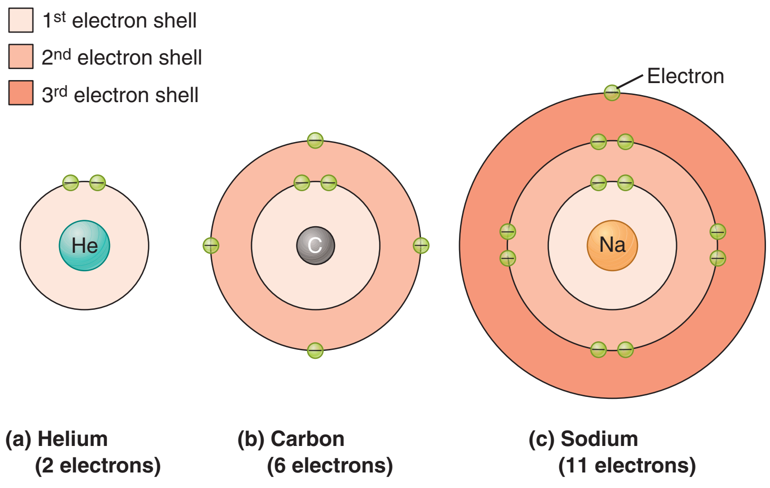
PET scan of the brain. The bright yellow areas use more glucose than normal and may indicate tumors.

But that's not the end of the good news about radioactivity. Once a thyroid tumor has been diagnosed, radioactive iodine can be used to treat the tumor. In this instance the patient can be injected with a larger dose of radioactive iodine. As the tumor accumulates radioactive iodine, the concentration of radioactivity becomes high enough to kill cancer cells. For some patients with thyroid cancer, this is the only treatment required.

Electrons Are Arranged in Electron Shells around the Atomic Nucleus

Recall that we compared atoms to a solar system: electrons follow particular paths about the nucleus, just as planets follow particular orbits around the sun. Just as some planets orbit nearer than others to the sun, some electrons orbit close to the nucleus and others orbit at various distances farther away. The paths along which the electrons move can be thought of as tracing out a set of hollow spheres—called *electron shells*—the smaller ones nested inside the larger ones like a child's set of nesting balls. The nucleus sits in the center of these spheres.

If we arrange atoms by size, we find that the two smallest—hydrogen and helium—have only one electron shell. As atoms grow larger—having nuclei with more neutrons and protons—additional electron shells are necessary to hold the electrons. Atoms can have up



(a) Helium (2 electrons) (b) Carbon (6 electrons) (c) Sodium (11 electrons)

Figure 2.4. Electron shells. Electrons are organized in energy levels called electron shells. **A.** Helium has a single shell. **B.** Carbon has two shells. **C.** Sodium has three shells. *Which of these atoms has a full outer shell?*

to five shells, each with a greater capacity than the previous one. The innermost (smallest) shell has a capacity of two electrons; that is, it can contain either one or two electrons. Again, hydrogen and helium are the only atoms with just one shell: hydrogen has one electron in the shell and helium has two (Fig. 2.4A). All other atoms have more electrons and, therefore, need more shells. The second shell has a capacity of eight electrons; the third shell has a larger capacity but is quite stable if it also contains eight electrons.

Two examples, using carbon and sodium, will illustrate the point. Carbon needs two shells to hold its six electrons: two electrons in the inner shell and four electrons in a second, outer shell (Fig. 2.4B). On the other hand, sodium, with 11 electrons, has 2 electrons in the inner shell (which fill it up), 8 in a middle shell (which fill it up), and 1 in a third shell, which leaves room for more electrons to join (Fig. 2.4C). Atoms are stable when they have eight electrons in the second shell and, if necessary, the third shell. Since both carbon and sodium have fewer than eight electrons in their outer shells, they are unstable and more readily participate in chemical reactions (see below).

Referring back to the periodic table (Fig. 2.2), the elements are arranged into rows based on the number of electron shells—hydrogen (H) and helium (He) have one shell, while potassium (K) and calcium (Ca) have four shells.

Electrically Charged Atoms Are Called Ions

Remember that protons are positively charged and electrons are negatively charged. If the number of protons and electrons in an atom is equal, the positive and negative forces exactly cancel one another and the atom has no *net* electrical charge. However, if the number of protons and electrons is not equal, the atom may have a *net* positive or negative electrical charge, in which case the atom is referred to as a positive or negative **ion.** Elements at the far left and right of the periodic table tend to form ions, while those closer to the middle do not.

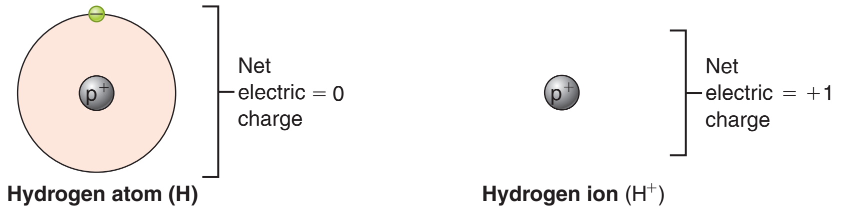
If electrons outnumber protons, the ion's net charge is negative, and if protons outnumber electrons, the ion's charge is positive. For example, a hydrogen atom missing its electron is a *hydrogen ion* (H^+, which is actually just the remaining proton) (Fig. 2.5A). A positively charged ion is called a **cation,** and a negatively charged ion is called an **anion.** If the number of protons outnumbers the number of electrons by two, as is the case with magnesium (Mg) ions, the magnesium atom has a net charge of +2, and the notation is written Mg^{++} or Mg^{2+}. Likewise, if electrons outnumber protons, as they do in negatively charged ions, the same principle applies. For example, an ion of the element chlorine (Cl) has one more electron than the number of protons and is written as Cl^-.

Ions are attracted to and tend to interact with other atoms that have an opposite charge. This characteristic of ions renders them very important in body processes—they can give us the power to move mountains! Ions will move a short distance to meet up with their opposing charge: this tiny movement enables muscle contractions that can lift a shovel full of dirt (or push the start button of a bulldozer).

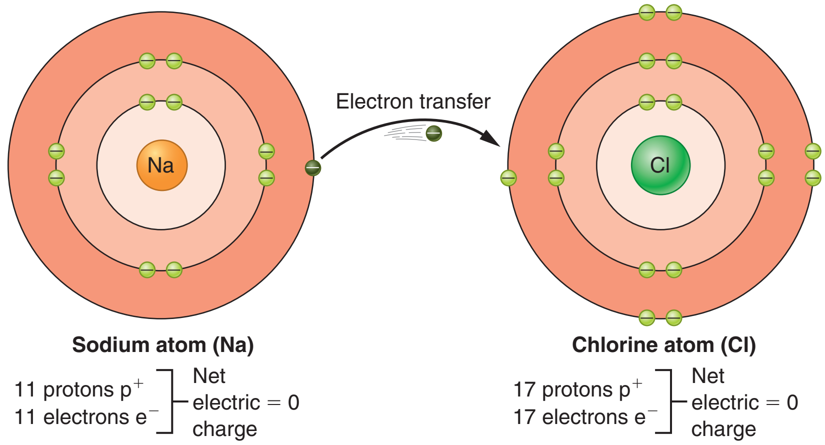
Case Note

2.3. Joe's treatment involved the administration of bicarbonate (HCO_3^-). Based on the negative sign in the formula, is bicarbonate an anion or a cation?

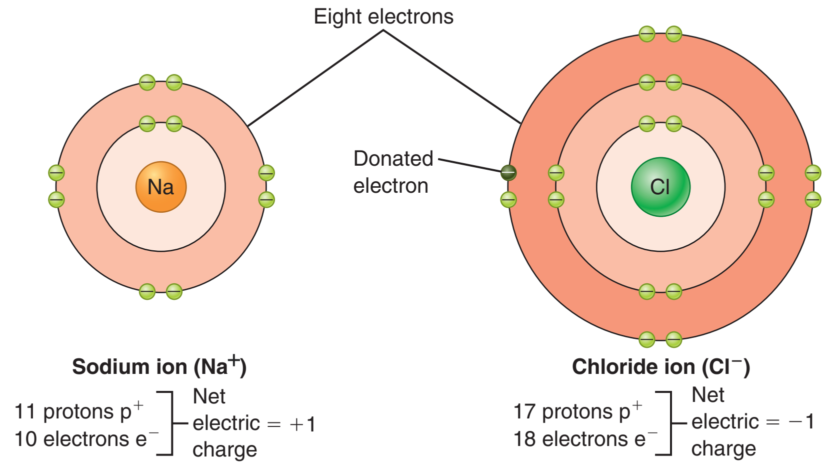
Atoms that have one or two electrons in their outermost shell are likely to *lose* them, which can empty the shell,



(a) The hydrogen atom and the hydrogen ion



(b) Sodium and chlorine atoms



(c) Sodium and chloride ions

Figure 2.5. Formation of ions. A. The hydrogen atom has a single electron in its outer shell. It loses this electron to another substance (not shown) to become a positively charged hydrogen ion (H^+). **B.** Sodium has one electron in its outer shell, and chlorine has seven. Sodium readily donates a single electron to chlorine. **C.** This reaction produces sodium (Na^+) and chloride (Cl^-) ions. Both ions now have eight electrons in their outer shell. *How many electrons are in the middle shell of the chlorine ion?*

leaving the underlying full shell as the outermost one and resulting in the formation of a positive ion. Atoms with six or seven electrons in the outermost shell tend to *gain* electrons to fill the shell, which stabilizes the shell (brings it near or to eight electrons) and forms a negative ion.

Remember This! **For a particular atom, changing the number of electrons results in an ion, changing the number of neutrons results in an isotope, and changing the number of protons results in a different element.**

Atoms gain or lose electrons by interacting with other atoms. Let's turn our attention back to the chlorine atom. Chlorine has only seven electrons in its third and outermost shell; so, in order to reach the stability of eight electrons, it is eager to "steal" an electron from another atom (Fig. 2.5B). On the other hand, sodium has only a single electron in its outer shell and is eager to lose it so that the underlying stable shell of eight electrons will become outermost. With the transfer of one electron from sodium to chlorine, both atoms attain a stable outermost electron shell containing eight electrons. Sodium is now a positively charged ion (Na^+) with 10 electrons

(2 in the innermost shell and 8 in the second, outermost shell) and 11 protons (Fig. 2.6C). And chlorine is now a negatively charged ion (Cl^-) with 18 electrons (2 in the innermost shell, 8 in the second shell, and a stable 8 in the third, outermost shell) and 17 protons. We refer to this form of chlorine by a different name—*chloride*—to signify the ionic form of the element.

Pop Quiz

2.4 Which subatomic particles do not have an electric charge?

2.5 The atomic number of neon is 10 and its atomic mass is 20. How many protons and neutrons are present in a neon atom?

2.6 True or false: Anions have more protons than electrons.

2.7 Aluminum atoms have 13 electrons. How many electrons will be in the outer shell of an aluminum atom?

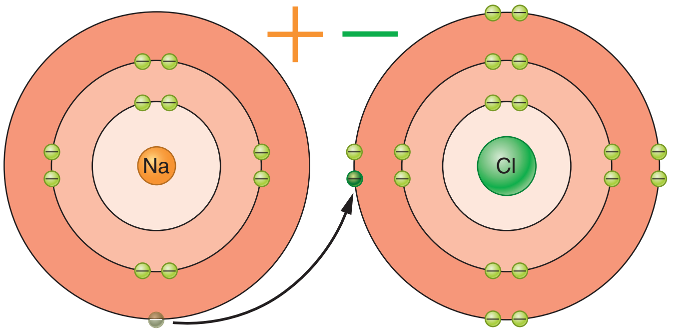
2.8 A dangerous form of oxygen is written as O22. Remember that oxygen has eight protons; how many electrons does this form of oxygen have?

Chemical Bonds

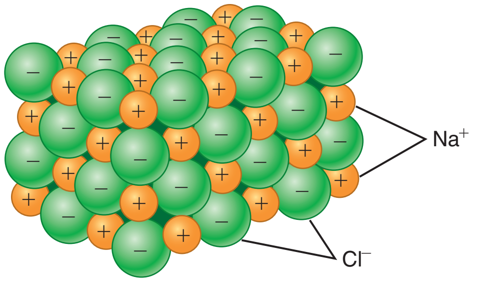
People do not live alone. Each person forms bonds of varying strength and duration with other people, ranging, say, from a 50-year marriage to a 5-minute conversation on the bus. In a similar way, atoms form bonds that are strong and durable or weak and transient.

Ionic Bonds Form between Ions

Remember that ions are formed when one atom *donates* an electron and another accepts it. However, just like a child who lends a treasured book to her brother, the electron donor stays close to the electron acceptor and does not completely lose control of the donated electron. Moreover, as a result of the donation, the receiving atom becomes negatively charged, the donor positively charged, and the two are strongly attracted to each other. The interaction between these opposing charges is called an *ionic bond*. For example, sodium (Na) transfers an electron to chlorine (Cl), forming Na^+ and Cl^- ions. The Na^+ and Cl^- ions form an ionic bond because of their opposing charges. The resulting substance is sodium chloride (NaCl, or table salt) (Fig. 2.6A).



(a) Na^+ and Cl^- attract each other, forming an ionic bond.



(b) A salt crystal results from ionic bonds between many Na^+ and Cl^- ions.

Figure 2.6. Ionic bonds. A. Positively charged sodium (Na^+) ions form ionic bonds with negatively charged chloride (Cl^-) ions. Note that both ions have stable outer electron shells. **B.** Many sodium and chloride ions are ionically bonded to form a salt (NaCl) crystal. Each ball represents a sodium (orange) or chloride (green) ion. *Which element loses a shell in the formation of NaCl, Na or Cl?*

NaCl is an example of a **compound,** a substance containing at least two *different* elements linked by a chemical bond. And any substance that is created by an ionic bond, not just NaCl, is a **salt.** Salts have a crystalline structure, created by arrays of alternating cations and anions interacting by ionic bonds (Fig. 2.6B).

Salts separate into their ionic parts when dissolved in water; that is, they *ionize*. For example, when NaCl is dissolved in water, it divides into Na^+ and Cl^- ions. Water containing ions conducts electricity easily, whereas pure water without dissolved ionic compounds does not. In fact, salts are also called **electrolytes** because they conduct electricity when dissolved in water.

Case Note

2.4. Joe was given an intravenous solution containing electrolytes. What is another name for electrolytes?

Molecules Are Atoms Joined by Covalent Bonds

Covalent bonds are bonds formed when atoms *share* electrons. Much like two children content with sharing a book, covalently bonded atoms share one or more electrons to fill their outer shell. The shared nature of covalent bonds renders them stronger than ionic bonds. For example, covalent bonds (unlike ionic bonds) remain intact when a covalently bonded compound is dissolved in water.

Two or more atoms linked by covalent bonds create a **molecule.** Chemical formulas are used to illustrate the structure of molecules. These formulas consist of symbols for the elements (H for hydrogen, C for carbon, and so on) together with a subscript number that indicates the number of atoms involved. For example, the symbols in the formula for water, H_2O, indicate two atoms of hydrogen combined with one atom of oxygen. And the formula for the main sugar in blood (*glucose*) is $C_6H_{12}O_6$, which indicates that a molecule of glucose is composed of 6 atoms of carbon, 12 atoms of hydrogen, and 6 atoms of oxygen.

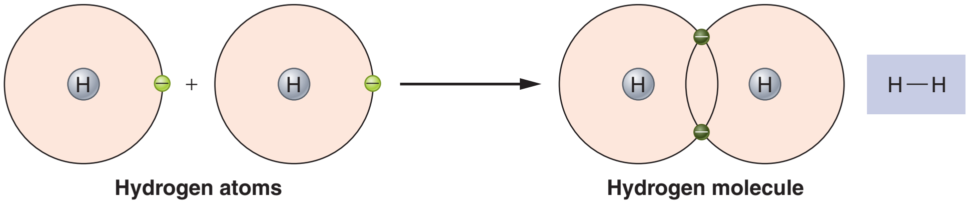
Remember This! **Two or more atoms linked together by covalent bonds form a molecule. The atoms can be the same, such as two hydrogens, or different, such as hydrogen and oxygen. Two or more atoms of different elements linked together by any chemical bond form a compound. Thus, H_2O is both a molecule and a compound.**

Even in the invisible world of atoms and molecules, form and function go together. Because the molecules we are studying are described on flat pieces of paper, it is easy to forget that molecules are three-dimensional objects; moreover, the function of a molecule depends not only on the atoms it contains but also on its three-dimensional shape (form). For example, glucose, which makes up part of table sugar, can assume two slightly different shapes: d-glucose and l-glucose. Even though these two forms contain identical sets of atoms, only one shape, d-glucose, can be burned for energy—l-glucose just doesn't "fit" the body's chemistry. Therefore, only d-glucose goes into sugar cookies.

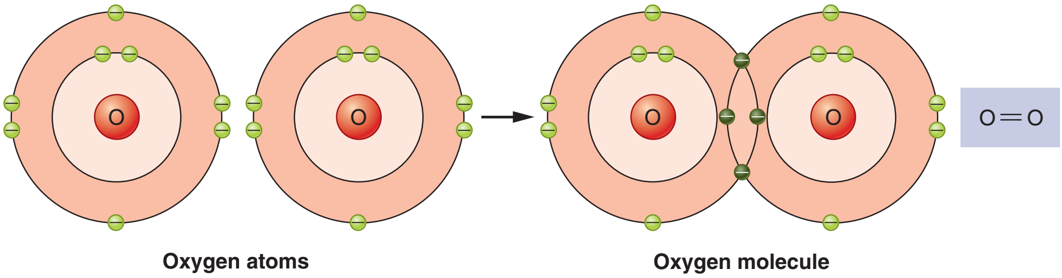
In Covalent Bonds, Atoms Share Electrons

The formation of hydrogen gas (H_2) from two hydrogen atoms is a good example of covalent bonding. Recall that an atom of hydrogen is formed of a nucleus that contains a single proton, which is orbited by a single electron (Fig. 2.7A, left side). When two hydrogen atoms (H) unite, the electron of each atom orbits *both* hydrogen nuclei, each electron flitting back and forth from one atom to the other (Fig. 2.7A, right side). This electron sharing is so complete that each atom of hydrogen acts as if its outer (and only) electron shell is filled to its stable capacity of two electrons.

In the above example, the covalent bond is composed of a single electron from each hydrogen atom. However, atoms can bond by each contributing two



(a) A single covalent bond



(b) A double covalent bond

Figure 2.7. Covalent bonds. A. Two hydrogen atoms form a single covalent bond when they share their single electrons. **B.** Two oxygen atoms form a double covalent bond by sharing two electrons each. *How many electrons are circling the hydrogen gas molecule?*

electrons instead of one, as when two oxygen atoms combine to form oxygen gas (O_2) (Fig. 2.7B). The resulting bond, which is quite sensibly called a *double bond*, is a very strong covalent bond, which stores a lot of energy.

In both of these examples, the electrons are equally shared between the two atoms. Figure 2.7 also provides a simple way of illustrating covalent bonds—drawing lines between the symbols for the bound atoms. A single line indicates that the atoms form a single bond (H–H in Fig. 2.7A); a double line indicates a double bond (O=O in Fig. 2.7B).

Polar Covalent Bonds Can Form Polar Molecules

We've said that covalent bonds can also form between atoms of different elements. Water, for instance, consists of one hydrogen atom and two oxygen atoms bound by covalent bonds (H_2O) (Fig. 2.8). However, different atoms do not share as nicely as identical atoms. Whichever one has a heavier nucleus (with more protons) exerts a stronger pull on the shared electrons than the atom with the lighter nucleus (fewer protons). The electrons, therefore, spend more time hanging out with the heavier atom. This causes the heavy end of the bond to be slightly negative and the lighter end to be slightly positive. Because the ends of the bond have opposing positive and negative charges, such a bond is called a *polar covalent bond*. That is, the bond has a negative pole and a positive pole, much like a magnet.

Nitrogen and oxygen, but not carbon, frequently form polar covalent bonds with hydrogen atoms. For instance, oxygen is heavier (has more protons) than hydrogen. Thus, when oxygen and hydrogen bond, the shared electron gravitates toward the oxygen, making that end of the bond more negative.

Because of the arrangement of its polar bonds, the water molecule as a whole is also polar. The oxygen side of the molecule is slightly negative and the hydrogen side slightly positive. The key word here is *slightly*: The positive and negative charges of polarized molecules are much weaker than the positive and negative charges associated with ions. For this reason, chemists designate these weak polarities with special symbols, δ^+ and δ^- (Fig. 2.8, right side).

Because a polar molecule has two poles—that is, regions of opposite charges separated by a short distance—it is also called a *dipole*. Molecules containing oxygen and nitrogen are frequently dipoles. By contrast, charges are evenly distributed in *nonpolar molecules*, which either are composed of a single element, such O_2, or, like most fats, contain only carbon and hydrogen.

Strong Covalent Bonds Create Functional Groups

Some covalent bonds are so strong that the participating atoms behave like a single atom, much like a close-knit group of friends. These atomic gangs are called *functional groups*. The hydroxyl functional group (OH), for instance, rarely dissociates into oxygen and hydrogen atoms. Five functional groups that play important roles in the chemistry of life are identified in Table 2.1. Functional groups form covalent bonds with other atoms and functional groups, creating large, complex

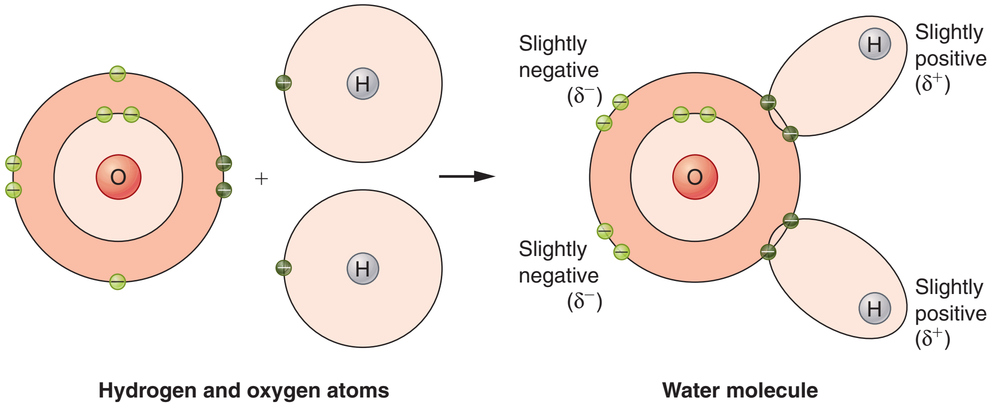
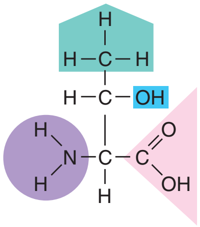
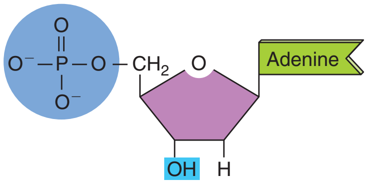
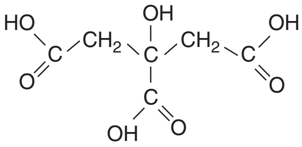


Figure 2.8. Polar covalent bonds. Water is formed by one oxygen atom and two hydrogen atoms bound by polar covalent bonds. The slightly positive end of the water molecule is shown by the symbol δ^+; the negative end by the symbol δ^-. *Which side of the water molecule is slightly positive: the hydrogen end or the oxygen end?*

Table 2.1 Functional Groups

FUNCTIONAL GROUP	FORMULA	EXAMPLES
Hydroxyl	–OH	 Threonine (amino acid)
Carboxyl	–COOH	
Amino	$-NH_2$	
Methyl	$-CH_3$	
Phosphate	PO_4^{-2}	 Adenine neucleotide (DNA)

molecules. Citric acid, the major component in lemon juice, is one such example. Its chemical structure is shown below.



Case Note

2.5. How many functional groups can you find in citric acid?

Some Molecules Are Ions

Recall that salts dissolve in water to form ions. Some salts contain more than two atoms; $CaPO_4$, for instance, is a salt that makes bones hard. When $CaPO_4$ dissolves in water, the PO_4 functional group behaves like a single atom. It receives two electrons from calcium, resulting in Ca^{2+} and PO_4^{2-}. The atoms in PO_4^{2-} are covalently bonded, forming a molecule, but the entire molecule has a negative charge. PO_4^{2-} is thus a *molecular ion* and can participate in ionic bonds with oppositely charged ions.

Highly polar molecules become ions by interacting with water molecules. In acetic acid (CH_3COOH), for instance, the H of the COOH end is virtually a naked proton because its electron spends all of its time at the other end of the molecule. When dissolved in water, this neglected proton forms a covalent bond with a water molecule, like a neglected cat bonding with a neighboring family, resulting in two ions, H_3O^+ and CH_3COO^-. Other molecules, such as ammonium (NH_3), perform the opposite feat, grabbing a hydrogen ion (i.e., a proton) from water to form OH^- and NH_4^+. As discussed later in the text, proton donors are called *acids* and proton grabbers are called *bases*.

Remember This! **In ions, the number of protons and electrons is unequal. Dipoles contain equal numbers of protons and electrons, but the electrons spend more time near some protons than others.**

Case Notes

2.6. A laboratory test measured the amount of three molecules in Joe's blood: bicarbonate (HCO_3^-), carbon dioxide (CO_2), and oxygen (O_2). Which of these molecules are compounds?

2.7. Which of these measured molecules would participate in an ionic bond: bicarbonate or oxygen?

Molecules Interact through Intermolecular Bonds

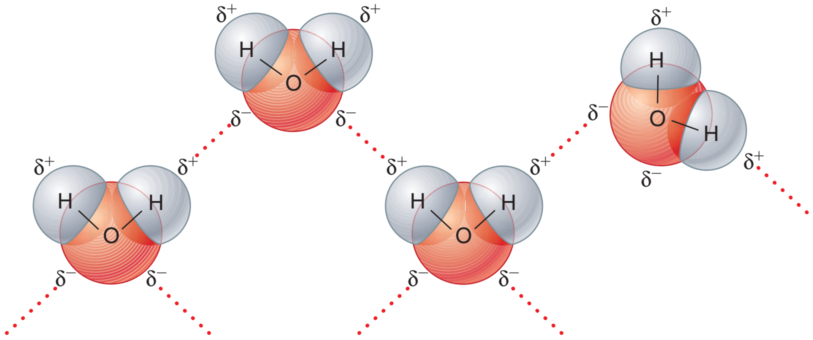
In the molecular world, opposites attract. For example:

- An ionic bond can form between a molecular cation (+) and a molecular anion (–).
- The negative end of a polar molecule can interact with a molecular cation, or the positive end of a polar molecule can interact with a molecular anion. Bonds between polar molecules and ions are called *dipole–ion bonds*.
- The negative end of a polar molecule can form a weak bond with the positive end of a different polar molecule, and vice versa. Interactions between two polar molecules are called *dipole–dipole bonds*.

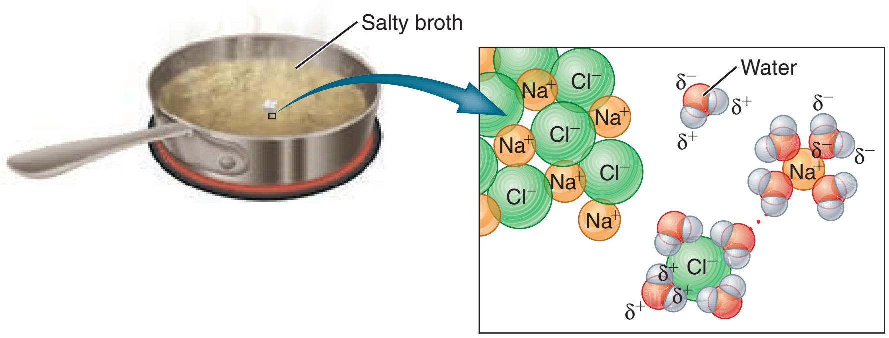
Now let's take a quick look at the most physiologically important type of dipole–dipole bond, the *hydrogen bond*.

Water Molecules Form Hydrogen Bonds

Hydrogen bonds are weak dipole–dipole bonds between the weak positive charge on the hydrogen atoms and the



(a) Hydrogen bonds between water molecules



(b) Salt dissociates in water

Figure 2.9. Intermolecular bonds. **A.** Hydrogen bonds form between polar water molecules. **B.** Table salt (NaCl) dissolves in water because polar water molecules form dipole–ion bonds with Na$^+$ and Cl$^-$. *What sort of bond is represented by the lines between the H and O symbols?*

weak negative charges of other polar molecules. In Figure 2.9A, for example, multiple water molecules form hydrogen bonds with one another. The negative oxygen end of one water molecule interacts with the positive hydrogen end of another water molecule.

Individual hydrogen bonds are much weaker than covalent or ionic bonds and are easily disrupted. Liquid water, for instance, is held together by hydrogen bonds. However, once water is heated above 100°F, the increased thermal energy of individual water molecules overcomes the hydrogen bonds and water evaporates. In contrast, when water reaches the freezing point, the bonds lock the sluggish water molecules solidly into place—and voila! Ice.

The important roles played by hydrogen bonds in the three-dimensional shape of proteins and DNA are discussed further on.

Charged Molecules Are Soluble in Water

The mixture of solute and solvent is called a **solution.** In a solution, the minor component is the **solute;** the major component is the **solvent.** *Solubility* is the ability of the solute to dissolve in the solvent.

The ability of charged molecules to form intermolecular bonds with each other determines their solubility. In the body, the most abundant polar solvent is water. Charged solutes are soluble in water because they can form dipole–dipole or dipole–ion bonds with water molecules. For instance, Figure 2.9B shows water forming dipole–ion bonds to dissolve NaCl into Na$^+$ and Cl$^-$. The slightly negative oxygen ends of water molecules interact with Na$^+$; the slightly positive hydrogen ends with Cl$^-$.

Polar and ionic molecules are said to be **hydrophilic,** or water-loving, because they are soluble in any watery substance, such as blood or cytosol. They are not soluble in nonpolar solvents because their charged solute molecules cannot form bonds with the nonpolar solvent molecules. For our purposes the most important nonpolar solvents are *lipids,* a class of molecules that includes oils and fats, which will be discussed later in the text.

Nonpolar molecules such as fats or oils cannot form dipole–dipole bonds with water and are thus not soluble in watery fluids like blood. Instead, they are soluble in other nonpolar substances and are called **hydrophobic,** or water-hating. In fact, it's the water that "hates" the lipid—water repels any nonpolar molecule even as it forms more bonds with other water molecules. This repulsion forces hydrophobic molecules to cluster together in little droplets, much as the icy Antarctic winds force penguins to huddle together. An everyday example of polar–nonpolar interactions can be found at the dinner table in a simple salad dressing of hydrophobic olive oil and hydrophilic vinegar. If left to stand for a few minutes, the olive oil and vinegar separate into two layers because the polar, watery vinegar solution drives away the nonpolar olive oil molecules.

Polarity and solubility are of immense practical importance in the administration of drugs–some drugs are polar and others are nonpolar. The polarity and thus the solubility of a drug determines how it can be administered. See the nearby Clinical Snapshot, titled *Drug Delivery Methods: Pill, Patch, Puff, Pump, or Poke?* to find out more about drug delivery methods.

Case Notes

2.8. A major solute in Joe's intravenous solution was bicarbonate ions. Do you think bicarbonate ions are hydrophilic or hydrophobic?

2.9. What sort of bond can bicarbonate ions form with water, if any?

2.10. Which solvent do you think was used for the intravenous solution: water or oil?

Remember This! **Like dissolves like. Polar solvents dissolve polar solutes, and nonpolar solvents dissolve nonpolar solutes.**

CLINICAL SNAPSHOT

Drug Delivery Methods: Pill, Patch, Puff, Pump, or Poke?

No one likes injections, but for most people with diabetes, multiple daily "pokes," or injections of insulin, remain a way of life. Insulin enables the body to metabolize glucose (blood sugar) for energy. Without glucose to burn for energy, human life is not possible. There are good reasons diabetic patients must inject insulin instead of swallowing it or smearing it on their skin. And the reasons are a matter of *chemistry*.

First, insulin cannot be effectively administered in a pill because stomach acid and intestinal digestive juices are designed to break down protein, and insulin, a protein, is destroyed by passage through the intestines. Second, insulin cannot be administered as a skin cream or patch because insulin, like most proteins, is hydrophilic (water-loving) and skin is hydrophobic (water-hating). On the other hand, hydrophobic drugs can be administered by placing them on skin, often as a patch; nicotine and estrogen are examples.

So, until recently, the only effective way to get insulin into the bloodstream has been by injection. Fortunately, advances in chemistry have recently led to some less painful methods of administration. Insulin molecules can be packaged into microscopic particles that can be inhaled deep into the lungs, where they can pass readily into the bloodstream. Another method is an insulin infusion pump, which can be worn in a small case like a cell phone and, periodically throughout the day, pumps insulin through a thin tube into a "port" in a blood vessel. What's more, under development is an artificial pancreas, a small device that senses blood sugar and releases exactly the right amount of insulin to control blood glucose. Soon, it seems, insulin injections will be a thing of the past for people with diabetes.

Insulin delivery. With the development of new insulin delivery methods, painful insulin injections may soon be a thing of the past.

Enzymes Promote Chemical Reactions

Life is maintained by an infinite number of chemical reactions, the sum of which is an organism's **metabolism.** Metabolic chemical reactions are usually described using symbols representing the chemical composition of molecules, which are linked by an arrow that indicates the way in which the reaction proceeds. For example, the following formula indicates how water (H_2O) and carbon dioxide (CO_2) combine to form *carbonic acid* (H_2CO_3):

$$CO_2 + H_2O \rightarrow H_2CO_3$$

Enzymes are specialized proteins that facilitate chemical reactions. Enzymes can be matchmakers, helping two molecules combine into one, or cleavers, chopping large molecules into smaller ones. Others are more subtle, adding an atom here or subtracting one there. Much like a kitchen knife is not changed after it has cut many carrots, an enzyme is not affected by the chemical reaction it facilitates and can enable many different reactions. Without the appropriate enzyme, the chemical reactions necessary for life would be so slow that life could not exist. The reaction above, for instance, is enabled by the enzyme *carbonic anhydrase*. The suffix *-ase* identifies most enzymes.

Recall that functional groups are particular atomic combinations that tend to stay together. Each functional group reacts in a characteristic way, usually with a functional group of a different kind on another molecule, regardless of the other parts of the molecules. Amino groups, for instance, usually react with carboxyl groups—a reaction important in the formation of proteins (see below).

Chemical Reactions Create or Break Chemical Bonds

Chemical bonds store energy, and it is not far-fetched to think of them as tiny sticks of dynamite. Just as dynamite is a store of energy capable of releasing an explosion of energy, foods are stores of energy that is released as their chemical bonds are broken in metabolic reactions. (Recall the box Basic Form, Basic Function: *Energy and Life* on page 35–36.)

For example, apples are sweet because the apple tree synthesizes *fructose,* or fruit sugar. The synthesis of fructose requires energy from the sun (as rays of light) and atoms of carbon, oxygen, and hydrogen (Fig. 2.10A). It gets these atoms in the form of molecules found in the air, water, and soil. For instance, water (H_2O) is a source of hydrogen and oxygen atoms and CO_2 a source of oxygen and carbon. Fructose thus stores the sun's energy in chemical form.

When we consume an apple, fructose enters our body cells. Enzymes break the chemical bonds joining the atoms, releasing the bond energy to power cellular activity. The sugar's atoms are reorganized into smaller molecules of water and carbon dioxide (Fig. 2.10B). In this way, the energy from sunlight captured and stored in sugar is released to power animal life.

According to whether they make or break chemical bonds, metabolic reactions can be classified in one of three ways:

- *Synthesis* reactions (Fig. 2.10A) are *anabolic* reactions that use chemical bonds to combine atoms or small molecules to produce new, larger molecules. Synthesis *consumes energy and stores it* and is essential in building cell and body parts. For example, growing children depend on synthesis reactions to build new bone and muscle, and these reactions are powered by energy obtained from food. Synthesis reactions can be characterized by the formula:

$$\text{Energy} + A + B \rightarrow AB$$

- *Decomposition* reactions (Fig. 2.10B) are *catabolic* reactions that break molecules into smaller parts. Decomposition releases stored energy and is essential in the maintenance of life. For example, growing children require energy to build new tissue as well as to run and jump. To obtain this energy, their cells decompose molecules of dietary nutrients into their component parts. Decomposition reactions can be characterized by the formula:

$$CD \rightarrow C + D + \text{energy}$$

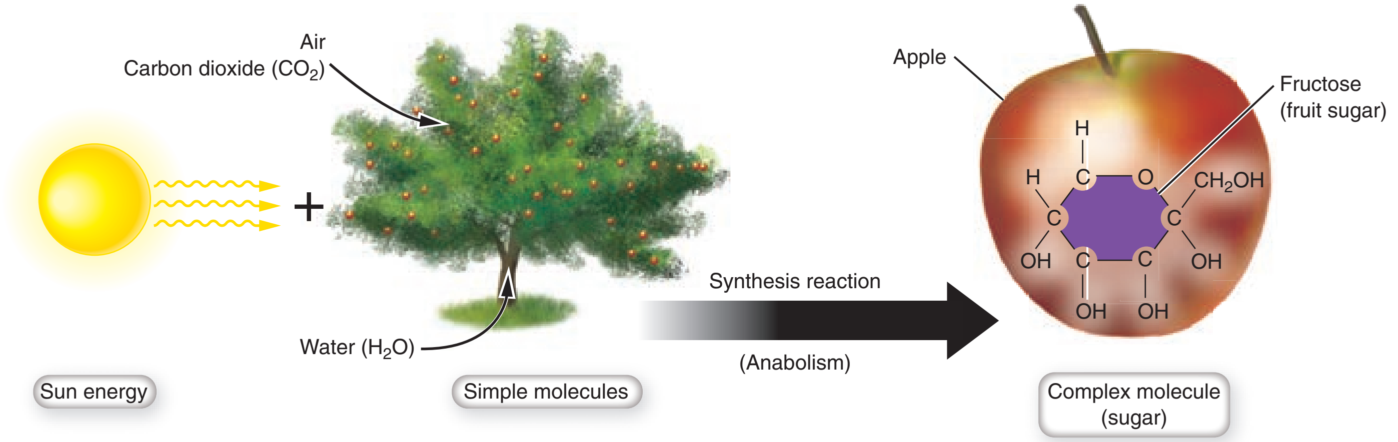
- *Exchange reactions* are metabolic reactions that involve both synthesis and decomposition, and chemical bonds are both made and broken. Energy production and consumption vary. In their simplest form, exchange reactions can be characterized by the formula:

$$EF + G \rightarrow EG + F$$

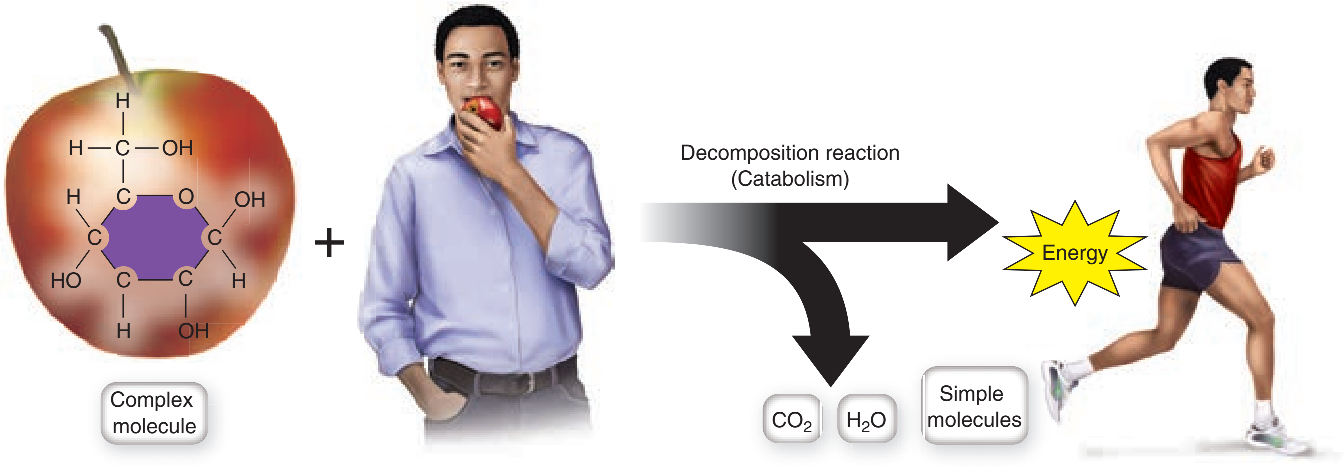
Remember This! **Anabolism is synthesis. Catabolism is breakdown.**

Chemical Reactions Can Proceed in Both Directions

Any chemical reaction can proceed in either direction; however, most reactions have a preferred direction that



(a) Synthesis reaction (Anabolism)



(b) Decomposition reaction (Catabolism)

Figure 2.10. **Chemical reactions. A.** Anabolic (synthesis) reactions combine atoms (or simple molecules) into larger ones, incorporating energy in the form of chemical bonds. **B.** Catabolic (decomposition) reactions release this energy by breaking the bonds. *Which type of reaction would be used to break a substance into its component parts?*

requires less energy, like pushing a stone down a slope. Reactions proceeding in the opposite direction are unusual because they require a great deal of energy, like rolling a stone uphill. However, some reactions can readily go back and forth, depending on the concentrations of the molecules involved. Such back-and-forth reactions are displayed by formulas that contain arrows pointing in both directions. Consider the reaction between hemoglobin (Hb, a molecule in red blood cells) and oxygen (O_2):

$$Hb + O_2 \underset{\text{tissues}}{\overset{\text{lungs}}{\rightleftharpoons}} HbO_2$$

An increase in the concentration of either hemoglobin (Hb) or oxygen will push the reaction toward the right, increasing the amount of HbO_2 formed. Conversely, a decrease in the concentration of either hemoglobin or oxygen will drive the reaction to the left, releasing oxygen from the hemoglobin. As explained in ➡ Chapter 13, this principle drives the reaction from left to right in the lungs, where oxygen concentration is high, and from right to left in tissues, where oxygen concentration is low because working cells have used up available oxygen.

Case Note

2.11. The concentration of oxygen was abnormally high in Joe's blood. Would he have more HbO_2 than usual, or less? Explain why.

Pop Quiz

2.9 Which of the following is a compound: H_2, H_2O, O_2?

2.10 True or false: Electrolytes are created by covalent bonds.

2.11 Is NaCl a molecule?

2.12 Which contains more hydrogen: a molecule of water or a molecule of glucose?

2.13 Which type of molecule contains an equal number of protons and electrons: a dipole or a molecular ion?

2.14 The glucose molecule is polar. Is it soluble in water or in fat? Explain why.

2.15 Hydrochloric acid (HCl) forms between the hydrogen cation and the chloride anion. Which type of bond holds HCl together—an ionic bond or a hydrogen bond?

2.16 Some women get injections of estrogen to prevent conception. Estrogen is a nonpolar molecule; should it be dissolved in water or oil?

2.17 Which atomic particles participate in chemical bonds?

The Chemistry of Living Things

To this point we have discussed processes and characteristics common to substances in all types of matter, from rocks to ears. Next we narrow our focus to the characteristics of substances found in the human body. These substances can be divided into two large categories: *inorganic* and *organic*. **Inorganic** substances are found in living and nonliving things, and contain no carbon atoms (an exception is carbon dioxide, CO_2). Examples include water (formed of hydrogen and oxygen) and table salt (formed of sodium and chloride). **Organic** substances (organic matter) contain carbon and originate in living things; that is, they are created by the chemical reactions (metabolism) of living things. Examples of organic molecules include carbohydrates, such as the starch found in wheat, and the proteins found in muscle.

Important Inorganic Substances Are Salts, Water, Acids, and Bases

The molecules and compounds of inorganic matter are smaller and less complex than organic ones. Remember that a salt, for instance, is a compound that contains a positively charged ion (cation) and a negatively charged ion (anion) bound by ionic bonds (Fig. 2.6). Examples include calcium phosphate ($CaPO_4$), which provides the strong framework of bones and teeth. Many salts perform their most important physiological functions in water after they dissociate into their component ions. Both calcium and sodium ions, for instance, are important in the transmission of nerve impulses ➡ (Chapter 8). Other important inorganic molecules include water, acids, and bases.

Water

Jupiter has its moons, Saturn its rings, and Mars its red hue, reflecting the abundance of iron in its soil and dust. Earth is called "the blue planet," a title that reflects its abundance of water. Life is not possible without *water*: It is the main ingredient in every cell, every tissue but bone, and every fluid in the body. For example, blood is about 90% water. The *physical properties* of water are important:

- It is a *solvent*. Recall that a solvent is a liquid in which a solute can be dissolved to form a solution. With the exception of fats and oils, most of the chemicals in living things are soluble in water, and with good reason: molecules cannot react unless they are in solution with one another. That is, the chemical reactions of life (metabolism) occur between molecules dissolved in water.
- It is a *lubricant*. Water is the main ingredient in fluids that smooth the movement of muscles, tendons, bones, and other tissues against one another.
- It is a *cushion*. Water insulates tissues and cells from the bumps and blows of daily life. For example, water is the main ingredient in the amniotic fluid in which the developing fetus floats in the mother's body; on a smaller scale, it is the main ingredient in the fluid (*interstitial fluid*) that surrounds every cell in the body.
- It is a *heat sink*. A heat sink is a substance that absorbs heat without much change in temperature. Metabolic reactions release large amounts of heat and body water absorbs it without much change; otherwise, body temperature would vary widely. This ability to absorb heat makes it easier for homeostatic mechanisms to maintain body temperature very close to the body's set point of 98.6°F.

Additionally, water is an important *chemical* in metabolic reactions. As discussed below, water participates in chemical reactions in one of two ways: It is either created or consumed.

Dehydration Synthesis

In some chemical reactions, water is *created* with the combination of a hydrogen ion (H^+) taken from one molecule and a hydroxyl ion (OH^-) from another molecule. This process is termed *dehydration synthesis*, because the two substances are "dehydrated" to create a new water molecule.

Our body cells use repeated cycles of dehydration synthesis to assemble the large molecules they need, such as new proteins. Called **polymers,** these large molecules are essentially long chains of similar or identical molecular subunits called **monomers,** which can be likened to beads strung into a necklace—the bead is the monomer; the necklace is the polymer.

The process of dehydration is analogous to assembling beads into a necklace. Imagine a partially assembled chain that has been capped to prevent the beads from falling off. Each loose bead similarly has a small protective cap. Removing the protective caps and sticking the new bead onto the chain lengthens the necklace.

In a similar fashion, monomers are added to short polymers to create longer polymers (Fig. 2.11A). Following removal of the OH and H "caps" on the polymer and monomer (respectively), the monomer is joined to the end of the short polymer. In doing so, this reaction stores energy in the bond between the newly joined pieces, and the OH^- and H^+ combine to form water. Dehydration synthesis is anabolic, because it creates a larger molecule from smaller ones.

Hydrolysis Reactions

Water is *required* and *consumed* for the breakdown of a polymer into shorter polymers or into monomers. Contrary to dehydration synthesis, this process is analogous to removing a bead from a necklace. In this process, a water molecule is inserted into a polymer, breaking its bonds. The water molecule separates into a hydrogen ion (H^+) and a hydroxyl ion (OH^-), and one part of the broken polymer receives the hydrogen ion (H^+) while the other part receives the hydroxyl ion (OH^-) (Fig. 2.11B). This reaction liberates energy stored in the chemical bond. This process is called **hydrolysis** because water is broken (*hydro-*, water; and *-lysis*, break) into its constituent parts.

Acids and Bases

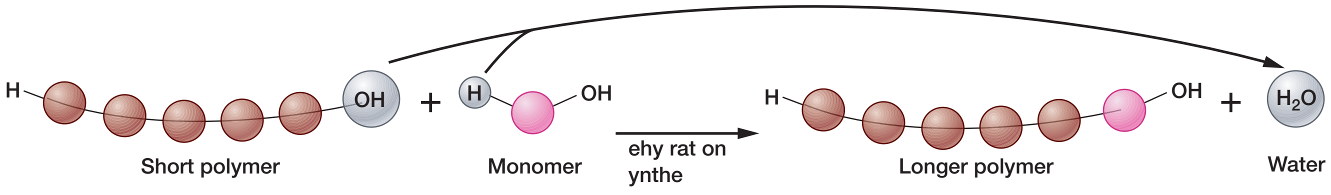
An **acid** is a chemical that releases (increases) hydrogen ions (H^+) when dissolved in water. Acids are important because they play an important role in metabolism: the hydrogen ions (H^+) they release are eager to interact with other chemicals. Acids that ionize completely and release large amounts of H^+ ion are strong acids; those that ionize only partially and release fewer H^+ ions are weak acids.

For example, *hydrochloric acid* (HCl), the chemical that gives stomach juice its acidity, is a strong acid because it ionizes completely:

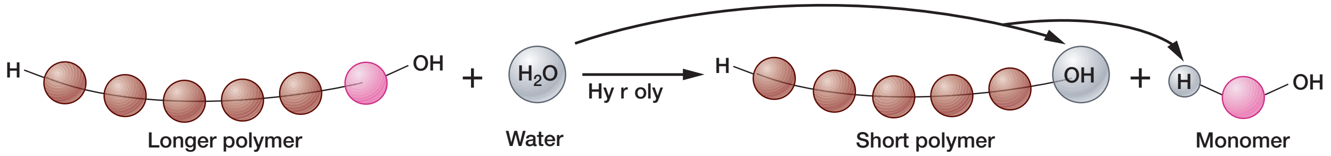
$$HCl \rightarrow H^+ + Cl^-$$

In contrast, *carbonic acid* is a weak acid created as cells that produce *carbon dioxide* (CO_2), a waste product of energy metabolism. The CO_2 mixes with water to create carbonic acid in the following way:

$$CO_2 + H_2O \rightarrow H_2CO_3$$



(a) Dehydration synthesis



(b) Hydrolysis

Figure 2.11. Water reactions. Dehydration synthesis generates a water molecule as it assembles polymers from monomers. Hydrolysis breaks a water molecule to remove monomers from a polymer. *During dehydration synthesis, is the monomer added to the hydroxy (OH) end or the hydrogen (H) end of the polymer?*

Carbonic acid (which is used in carbonated beverages to give them their tart taste and fizz) is weak because it does not completely ionize, as indicated by the following formula:

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

That is, only some of the H_2CO_3 molecules will dissociate, while other molecules will remain intact.

A **base** (*alkali, caustic*) is a chemical that decreases (absorbs) hydrogen ions when dissolved in water. The strongest bases release hydroxyl (OH^-) groups, which combine with H^+ ions to produce water (H_2O). For example*, sodium hydroxide* (NaOH, the main ingredient in household liquid treatments for clogged drains) is a strong base:

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

Weak bases, like weaker acids, do not completely dissociate in water and do not attract hydrogen ions with OH^-. The *bicarbonate ion* (HCO_3^-), the most abundant base in blood, is a weak base that attracts hydrogen with the HCO_3^- ion instead of the OH^- ion.

Case Note

2.12. The lemon juice Joe consumed was rich in citric acid. Citric acid does not completely ionize in water. Is it a weak acid or a strong acid?

The pH Scale Is a Measure of Acidity

A convenient way of measuring the strength of an acid or base is the **pH** scale, a series of numbers from 0 to 14, in which 0 is purely acidic (all H^+ ions) and 14 is purely basic (no H^+ ions, all OH^- ions). As the concentration of H^+ ions rises, the concentration of OH^- ions falls, and the solution becomes more acidic. As the concentration of OH^- ions rises, the concentration of H^+ ions falls, and the solution becomes more basic. Water is neither acidic nor basic; it is neutral and has a pH of 7.0. That is, it has an equal concentration of H^+ and OH^- ions.

Small changes in pH represent large changes in H^+ and OH^- concentration. A change of pH by one unit on the scale represents a tenfold (10×) change in the concentration of hydrogen (or hydroxyl) ions. For example, the pH of milk of magnesia (a laxative) is 10.5, and the pH of household ammonia is 11.5. This means that the ammonia has 10 times more hydroxyl ions (and 1/10 as many hydrogen ions) as milk of magnesia. The pH value of other fluids and everyday items is depicted in Figure 2.12. Notice that human blood is slightly basic, about pH 7.4, which reflects its concentration of bicarbonate (HCO_3^-) ions. As discussed immediately below, bicarbonate is an abundant base whose job is to absorb the acidic hydrogen-ion waste products of the body's energy metabolism.

Case Note

2.13. The pH of Joe's blood was 7.26. Does his blood contain more hydrogen ions or fewer hydrogen ions than normal?

Buffers Protect against Changes in pH

Human cells require a pH near 7.4; variation of just a few tenths of a unit can be life-threatening. To maintain pH at this homeostatic set point, the body utilizes *buffers* that resist changes away from 7.4 in either direction. Specifically, a **buffer** is a chemical that converts a stronger acid into a weaker acid (or a stronger base into a weaker base) that can be safely expelled from the body. Because energy metabolism produces an excess of acid, the body needs a way to neutralize hydrogen ions. One chemical that fills this role is the *bicarbonate* ion (HCO_3^-), which we encountered above. As hydrogen ions (H^+) are produced by metabolic reactions, bicarbonate neutralizes them in the following way:

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

Normally, the CO_2 produced by the reaction is blown off by the lungs ➡ (Chapter 13) and the small amount of water created by the reaction is absorbed into body water.

As the arrows in the formula indicate, the chemical reaction can go either way, absorbing or producing H^+ ions as needed to keep blood pH stable near 7.4. Repeated vomiting, for example, can cause the loss of gastric acid, resulting in a shortage of H^+ ions in the body. In such a case the reaction above would run "backward" (right to left) in order to produce more blood H^+ ions to make up for the loss of stomach acid.

Case Note

2.14. Joe's blood was analyzed for bicarbonate, oxygen, and carbon dioxide. Which of these molecules is a chemical buffer?

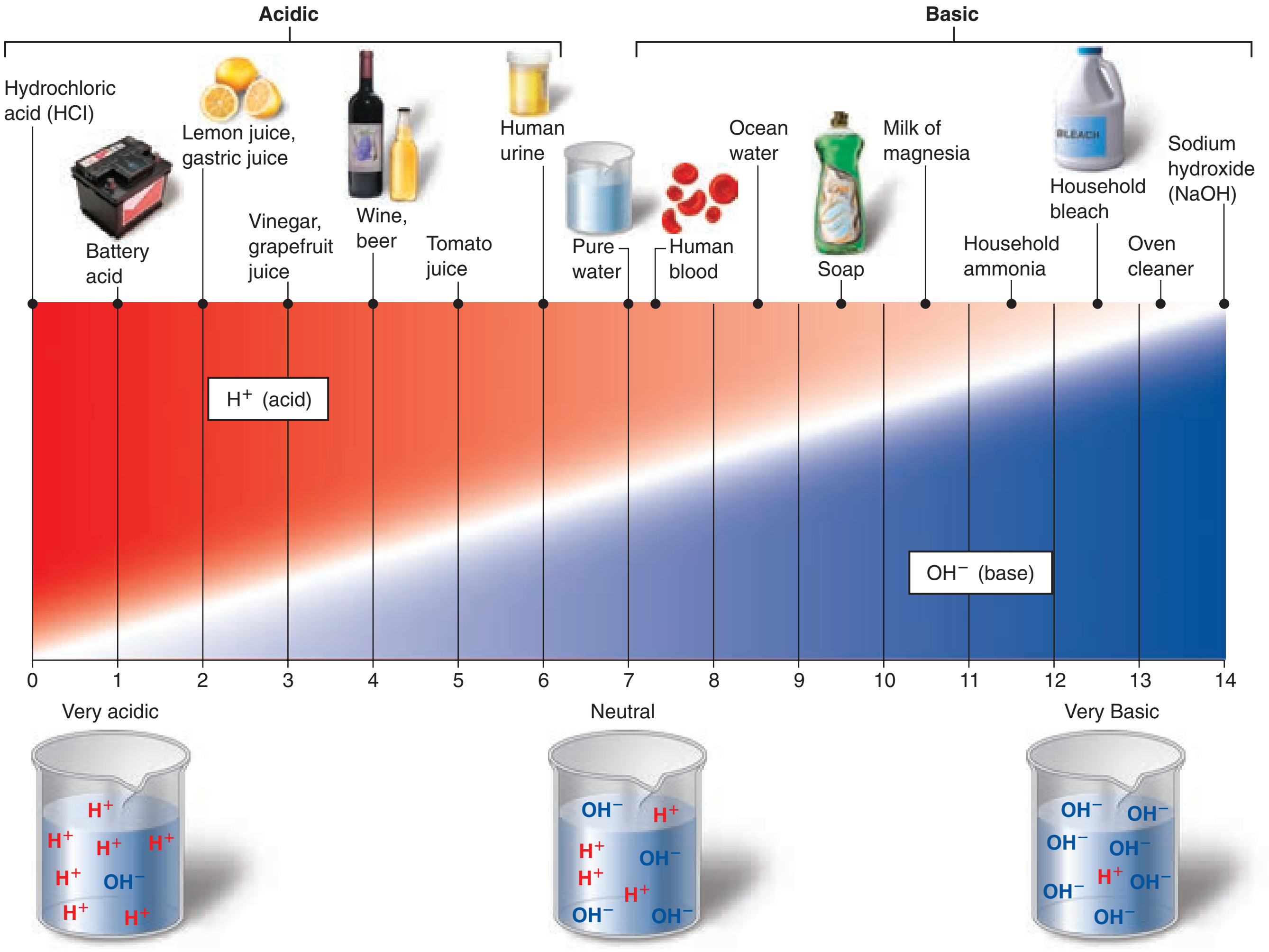


Figure 2.12. **The pH scale.** *What is the approximate pH of wine?*

Case Discussion

Chemistry in Context: The Case of Joe G.

Recall that Joe was sick because he was forced to drink a large amount of lemon juice, which contains a lot of acid.

The Problem: Excess Acidity. Joe had an acid–base problem (Fig. 2.13). In a short time he ingested far, far more lemon juice, a fairly strong acid, than the body could handle safely. In short, although lemon juice is a necessary ingredient in lemonade, it can be *toxic* (poisonous) in large amounts. Strictly defined, a **toxin** is a substance of plant or animal origin that causes disease when present at low concentration in the body. However, *any* substance, even water, can be injurious (toxic) in unusually large amounts, as was lemon juice in this instance. Refer to ➡ Chapter 16 for a discussion of water intoxication (poisoning).

Lemon juice is very acidic (pH 2); that is, it has a much higher concentration of hydrogen ions (H^+) than hydroxyl ions (OH^-). Joe's intestine absorbed these H^+ ions into his blood, so that it became more acidic than normal: the pH of his blood fell to 7.26 instead of remaining at 7.4, where it should be. Blood pH this low is not compatible with good health—Joe is breathing unusually fast and his brain function is affected: he is drowsy and nauseated.

The Solution (Part 1): Homeostatic Mechanisms. Joe must get rid of the excess acid he has consumed, which can be accomplished in one of two ways: through the kidneys, which work relatively slowly, or the lungs, which work much more rapidly. However, Joe cannot simply exhale or urinate pure acid. Instead, his body

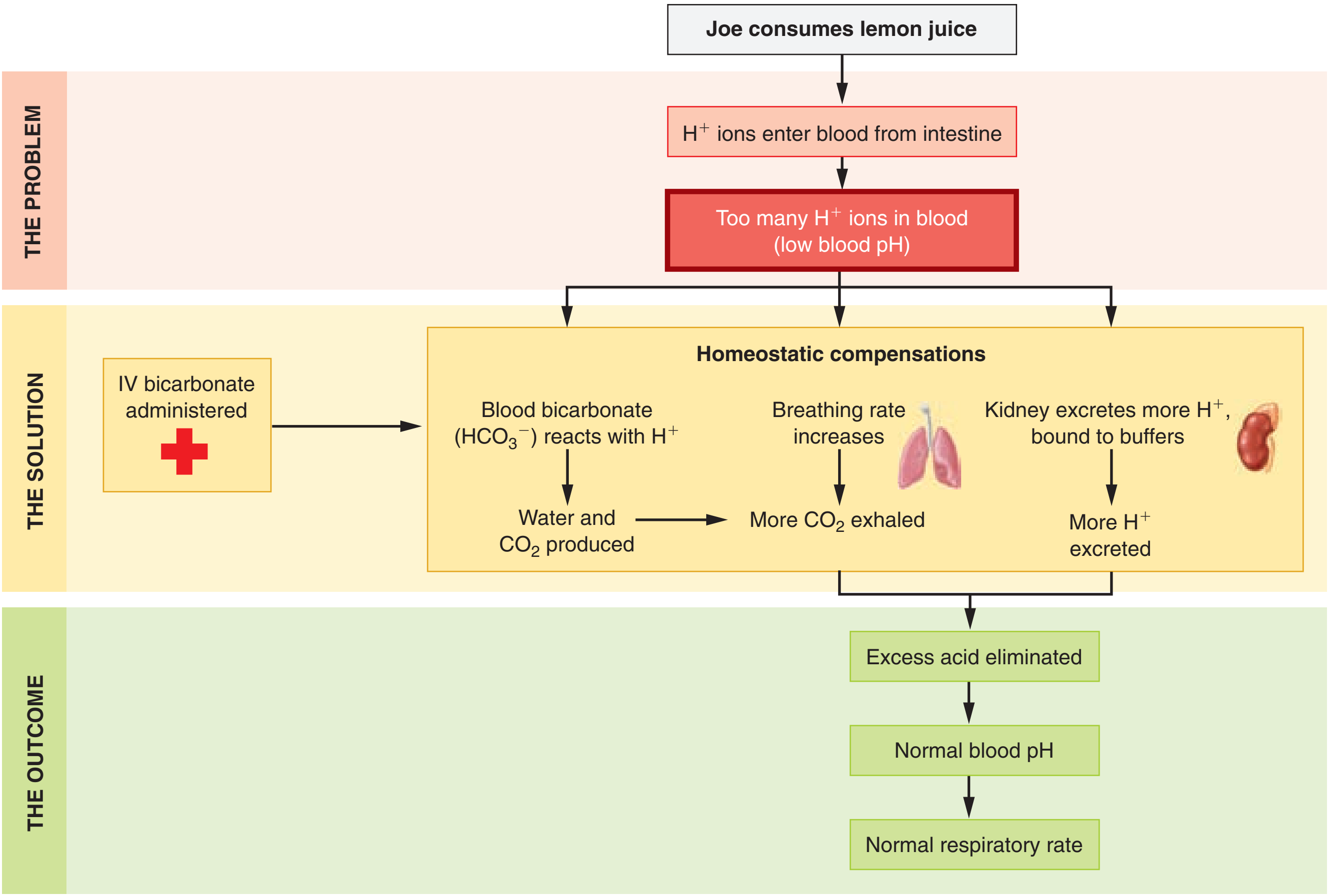


Figure 2.13. **Homeostasis and Joe G.** *Which organs help return Joe's blood pH to normal?*

must combine the acid with some other substance to convert it into a chemical that can be managed safely. Buffers to the rescue!

Buffers convert the hydrogen ions into compounds that can be safely excreted or exhaled. The excess H^+ ions in Joe's blood combine with blood bicarbonate (HCO_3^-), which is then converted into CO_2, which can be exhaled. The reaction is depicted below:

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

This reaction can proceed in either direction, but the high concentration of H^+ in Joe's blood drives the reaction to the right. In parallel, the hydrogen ions cause the brain to increase Joe's breathing rate so that the extra CO_2 can be exhaled rapidly. As Joe's lungs eliminate the excess acid as CO_2, Joe's blood pH rises and his breathing rate slows.

In the meantime, H^+ ions also combine with other buffers in the blood and kidneys ➡ (Chapter 16).

The Solution (Part 2): Medically Assisted Homeostasis. Unfortunately, Joe's symptoms suggest that his blood does not contain enough buffers to neutralize all of the hydrogen ions he has consumed, so the physician gave him extra bicarbonate buffer intravenously to assist in neutralizing and eliminating them. The success of the treatment and his homeostatic compensations is shown by Joe's reduced respiratory rate and normal blood biochemistry the following morning.

Case Notes

2.15. Bicarbonate decreases the concentration of hydrogen ions when it is dissolved in water. Is bicarbonate an acid or a base?

2.16. Look at the chemical reaction above. When Joe was given bicarbonate, did the reaction proceed to the left or to the right?

2.17. **When Joe was given bicarbonate, what happened to his blood H^+ concentration and his blood pH?**

2.18. **Which ion was responsible for Joe's fast breathing when he arrived at the hospital?**

2.19. **Would you consider rapid breathing to be a homeostatic mechanism in Joe's case? Explain.**

Important Organic Molecules Are Carbohydrates, Lipids, Proteins, and Nucleotides

Organic molecules are created only by living things but can be found in nonliving things. Bread, for example, is made of the organic matter synthesized by wheat. In the main, organic molecules are formed from atoms of carbon (C), hydrogen (H), oxygen (O), and nitrogen (N).

There are four types of organic molecules: *carbohydrates*, *lipids*, *proteins*, and *nucleotides*. These molecules are generally quite large and complex and are frequently described as *macromolecules*.

Carbohydrates

Carbohydrates (Fig. 2.14) are organic molecules composed of atoms of carbon, hydrogen, and oxygen arranged in short or long chains or rings. Carbohydrates are an important source of energy.

The simplest carbohydrates contain a single unit consisting of three to seven carbon atoms. Not surprisingly, such sugars are called *simple sugars* (or *monosaccharides*, from Latin *saccharum*, "sugar"). For example, glucose (Fig. 2.14A), the main sugar in blood, most often occurs as a six-sided ring (hexamer). The formula is $C_6H_{12}O_6$. Glucose, like many organic molecules, is often illustrated in a simplified form, as in the left-hand depiction in the figure. Each line represents a covalent bond, with a carbon atom at either end of the bond. Hydrogen atoms are not shown.

Monosaccharides are monomers: more complex carbohydrates are polymers composed of monosaccharides strung together somewhat like the links in a metal chain. The bonds between the links are forged by dehydration synthesis reactions (Fig. 2.11).

The combination of two simple sugars is a *disaccharide* (Fig. 2.14B); for example, the combination of two

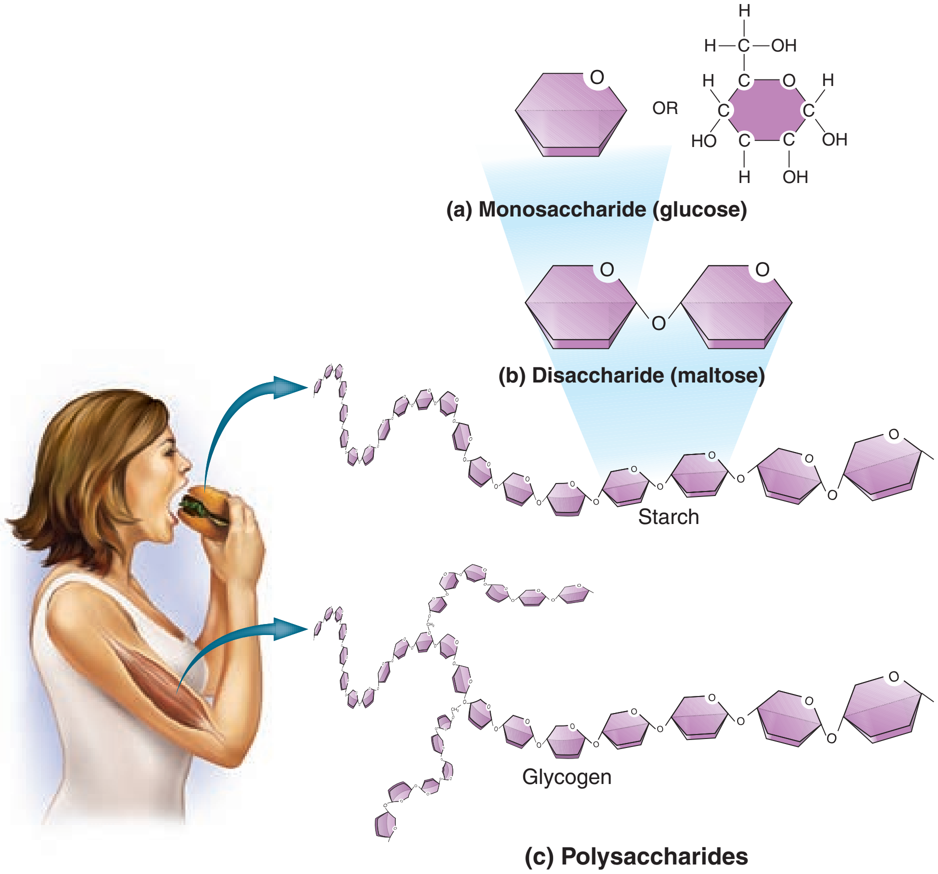


Figure 2.14. Carbohydrates. **A.** Monosaccharides consist of a single sugar molecule. **B.** Two monosaccharides combine to form a disaccharide. **C.** Many monosaccharides combine to form polysaccharides such as glycogen or starch. Glycogen is used to store carbohydrate in muscle and liver tissue. *What is the difference between glycogen and starch?*

glucose molecules produces the disaccharide *maltose*, and glucose plus fructose (another simple sugar, the primary sugar in fruits and honey) yields a disaccharide called *sucrose* (table sugar). The other physiologically relevant mono- and disaccharides are summarized in Figure 14.1 in ➡ Chapter 14.

A polymer composed of more than two monosaccharides is called a *polysaccharide*. These long, branched chains of simple sugars serve as a store of energy in plants and animals. In plants, the energy-storing polysaccharide is called **starch;** it is formed of glucose molecules linked together in a long chain. Starch is abundant in wheat, beans, and potatoes. In animals, the energy-storing polysaccharide is **glycogen;** it is also formed of long chains of glucose but is more highly branched than starch. Glycogen is stored in muscle and liver tissues.

Carbohydrates provide a ready source of energy for all physiological activities. However, disaccharides and polysaccharides are too large to enter cells and cannot be burned for energy until they are broken down into monosaccharides by hydrolysis reactions. Enzymes in the small intestine's lumen and wall break larger carbohydrates into monosaccharides, which are small enough to be absorbed by intestinal epithelial cells. Monosaccharides (especially glucose) are used for energy, in a reaction where each molecule of glucose is combined with six oxygen molecules. This reaction produces carbon dioxide, water, and energy, as shown in the following formula:

$$C_6H_{12}O_6 + 6\ O_2 \rightarrow 6\ CO_2 + 6\ H_2O + \text{energy}$$

Excess glucose not required to fuel cell activities is assembled into *glycogen* for storage in liver or muscle, or it is converted into *fat*.

A very small amount of carbohydrate is used to build cell structures. For example, some sugar molecules are used to build nucleotides (discussed below), which in turn are used to construct the DNA of our genes. Other carbohydrate molecules are combined with protein or lipid molecules to make glycoproteins or glycolipids, respectively. These important molecules are discussed later in the text.

Case Note

2.20. The most abundant disaccharide in lemon juice is sucrose. How many sugar units does sucrose contain?

Lipids

Lipids are slick, greasy, nonpolar substances that are hydrophobic. Solid lipids are frequently described as *fats;* liquid lipids as *oils*. They are composed mainly of atoms of carbon and hydrogen, with only a few oxygen atoms. For example, one common lipid has the formula $C_{57}H_{110}O_6$. Unlike carbohydrates, all of which have a similar chemical structure, lipids vary greatly in their *chemical* structure and are alike mainly in their *physical* characteristics: they are oily and not soluble in water. Some lipids important to human physiology are identified in Table 2.2.

Common Lipids Contain Glycerol and Fatty Acids

Despite their dissimilarity, many lipids contain two common building blocks: a molecule called **glycerol** and one or more *fatty acids* (Fig. 2.15A). *Glycerol* is an alcohol (it contains an OH group). A **fatty acid** is a long chain of carbon atoms (and their attached hydrogens) with an acidic molecule (the carboxyl group, COOH) at one end.

If all of the carbon atoms in a fatty acid chain are linked by single bonds, each carbon atom will have two hydrogen atoms attached to it. Such a fatty acid (and the lipids and storage fat made from it) is said to be *saturated* (with hydrogen). These fatty acids are quite straight; thus, any fat containing mostly saturated fatty acids will contain tightly packed, parallel chains of fatty acids and will be rather dense and solid at room temperature. Butter, lard, and the visible fat in meats are examples of saturated fats.

In contrast, if *any* carbon atoms in a fatty acid chain are linked by double bonds, these carbons will have only one hydrogen atom attached instead of two. Such fatty acids are said to be *unsaturated;* that is, they have less than the maximum possible number of hydrogen atoms. The double bond forms a bend in the fatty acid chain, so the fatty acids cannot be packed into neat parallel rows. Unsaturated fats, such as olive and corn oils, are thus less dense and tend to be liquid instead of solid at room temperature. Monounsaturated fatty acids have one double bond in the chain, and polyunsaturated fatty acids have more than one double bond.

The distinction between saturated and unsaturated fats also has important health implications: diets high in saturated fats promote harmful deposits of fat in blood vessels (atherosclerosis), a condition associated with high blood pressure, heart disease, and other diseases. The different types of dietary fat are discussed in greater detail in ➡ Chapter 14.

Table 2.2 Examples of Common Lipids

Name	Structure	Location	Function(s)
Triglycerides	Glycerol + 3 fatty acids	Primary lipid in fat tissue	Store energy; cushion bones; keep organs warm and protected
Phospholipids	Glycerol + 2 fatty acids + phosphate	Cell membrane	Make up cell membrane
Sphingolipids	Sphingosine + fatty acid + phosphate + amino acid	Cell membrane	Cell recognition
Cholesterol	Steroid	Cell membrane, liver, blood	Contributes to cell membrane; precursor to steroid hormones; component of bile
Vitamin D	Steroid	Formed in skin	Necessary for bone health
Sex hormones	Steroid	Formed in testes, ovaries, adrenal glands	Necessary for normal sexual and reproductive function
Corticosteroids	Steroid	Formed in adrenal glands	Regulate aspects of growth and metabolism
Aldosterone	Steroid	Formed in adrenal glands	Regulates salt and water balance
Prostaglandins	Derived from arachidonic acid (a fatty acid)	Formed everywhere	Many diverse functions; involved in pain and inflammation
Leukotrienes	Derived from arachidonic acid (a fatty acid)	White blood cells	Assist in inflammatory response
Thromboxanes	Derived from arachidonic acid (a fatty acid)	Produced by platelets	Involved in blood clotting
Fat-soluble vitamins (A, E, K)	Derived from isoprene (a small molecule containing C and H)	Obtained in the diet	Necessary for vision (A), blood clotting (K); vitamin E is an antioxidant

Case Note

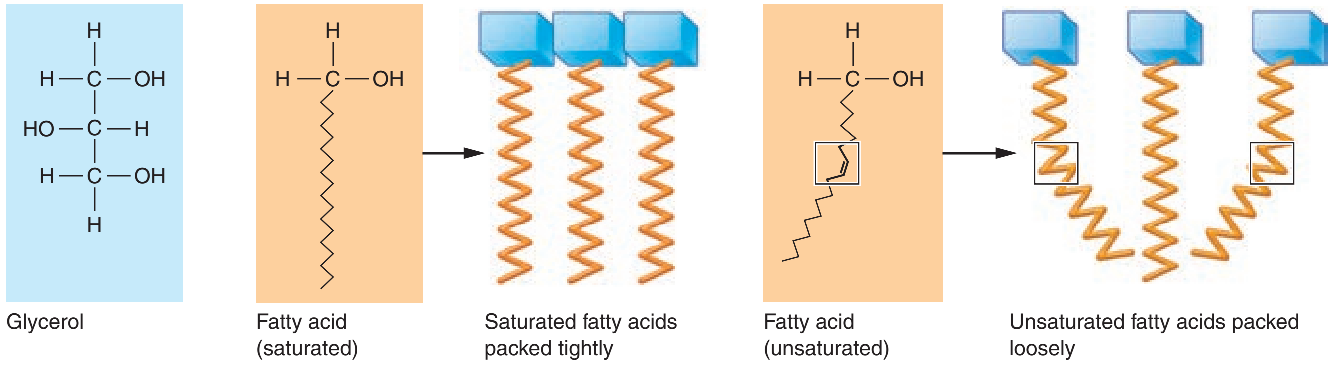
2.21. The most abundant fatty acid in lemon juice is called linoleic acid. It contains two double bonds. Is this fatty acid saturated or unsaturated?

Triglycerides Are Common Lipids in Foods and in the Body

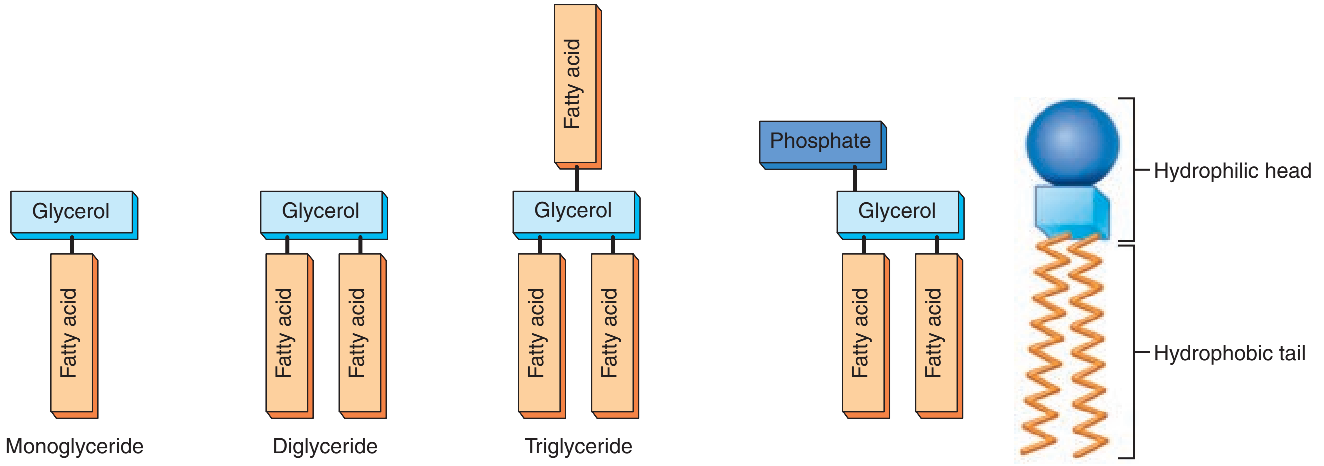
Glycerol and fatty acids can be combined by dehydration synthesis (Fig. 2.11) to form common lipids. Depending on the number of fatty acids attached to the glycerol molecule, this reaction will result in a monoglyceride (one fatty acid), diglyceride (two fatty acids), or **triglyceride** (three fatty acids) (Fig. 2.15B). A triglyceride is a type of lipid that is stored efficiently by the body to be burned later for energy. Triglycerides are the most space-efficient energy-storage molecules because they pack the most potential energy into the smallest space. Triglycerides contain nine calories per gram; carbohydrates contain only four. A **fat** is a lipid that is solid at room temperature. Triglycerides are commonly described as fats, because they are stored in the body in their solid form. Fats do not include phospholipids or steroids, two other lipids discussed next.

Case Note

2.22. The fatty acid limonine is responsible for the odor of lemons and lemon juice. Would this compound be hydrophilic or hydrophobic?



(a) Lipid building blocks



(b) Common lipids

(c) Phospholipids

Figure 2.15. Common lipids. A. Common building blocks for lipids include glycerol (light blue squares) and fatty acids (orange lines). **B.** These components can be built into mono-, di-, or triglycerides. **C.** They can also combine with phosphate (blue circle) to form phospholipids. *How many fatty acids are found in triglycerides?*

Phospholipids Are Common Lipids That Contain Phosphate

Phospholipids (Fig. 2.15C) are similar to triglycerides. Each is formed of a molecule of glycerol to which there are attached three other molecules: two fatty acid molecules (not three, as in triglycerides) and a phosphate group. Phospholipids have a unique quality: one end of the molecule (the phosphate end) is polarized and is, therefore, hydrophilic (soluble in water). The other end, the end containing the fatty acids, is not polarized and is not soluble in water (hydrophobic), but it is soluble in lipid.

It is true that oil and water don't mix, but where oil and water meet, the hydrophobic end of a phospholipid molecule can mix with oil and the hydrophilic end can mix with water. This quality is very important in everyday life and in the body. Dish detergent, for example, consists of molecules similar to phospholipids. Since one end of the soap molecule is soluble in water and the other in grease, the addition of soap molecules to water allows us to wash a sink full of greasy dishes. In the body, a phospholipid-containing chemical called *bile* acts in much the same way to help break down the fats in our foods.

Steroids Are Ring-Shaped Lipids

Steroids (Fig. 2.16) are similar to triglycerides and phospholipids in that they, too, are formed mainly of carbon and hydrogen and are slick, oily substances. However, steroids are structurally very different: they consist of four interlocking rings of carbon and hydrogen (the steroid nucleus) to which are attached short chains of other molecules. It is these attached short chains that govern the specific function of each steroid.

Cholesterol is one of the most important steroids in the human body. It is synthesized by the liver and need not

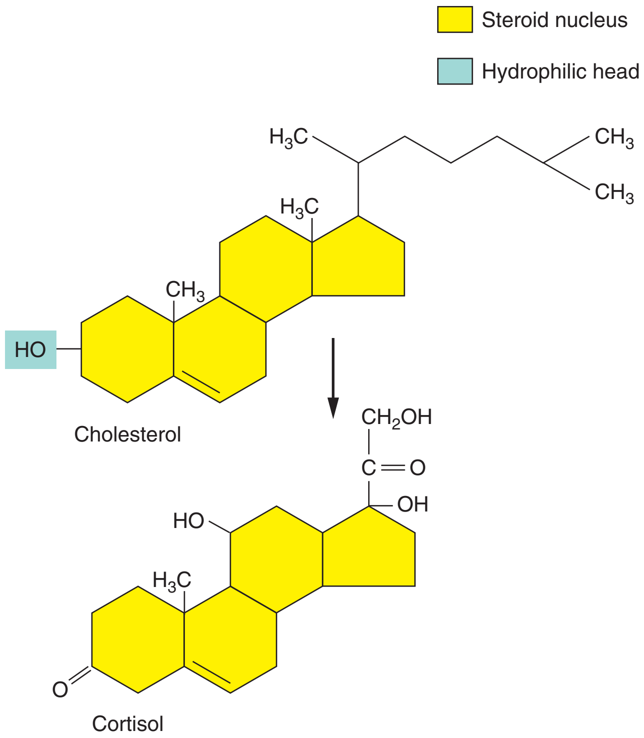


Figure 2.16. Steroids are ring-shaped lipids. Cholesterol contains the steroid nucleus of four carbon rings and a hydrophilic head group. Cholesterol can be converted into other steroids, such as cortisol, without modifying this steroid nucleus. *Which atom replaces the hydroxyl group as cholesterol is converted into cortisol?*

be consumed in the diet. However, it is present in foods of animal origin such as meat and cheese. Cholesterol is the basic building block for many other steroids called *hormones*. As discussed in ➡ Chapter 4, hormones act as signaling molecules that travel from one place to another and act as messengers in the regulation of body processes. For example, the steroid hormone *cortisol* helps regulate blood glucose levels.

Cholesterol, like phospholipids, contains hydrophilic and hydrophobic components. Cholesterol and phospholipids are the primary building blocks of the cell membrane, which is discussed in detail in ➡ (Chapter 3).

Prostaglandins Are Modified Fatty Acid Signaling Molecules

Prostaglandins are hydrophobic signaling molecules synthesized from fatty acids (Fig. 2.17). Their effects are both powerful and diverse. For instance, they cause blood vessels to dilate, are important in the functioning of the male and female genitalia, sensitize nerves to pain, and assist in regulating inflammation.

Nonsteroidal anti-inflammatory drugs (NSAIDs) such as ibuprofen treat pain by blocking the nerve-sensitizing effects of prostaglandins.

Proteins

Proteins comprise over half of the organic matter in the body. For example, skeletal muscle is mainly protein. Like carbohydrates and lipids, proteins are mainly composed of atoms of carbon, hydrogen, and oxygen; however, proteins also contain atoms of *nitrogen* (N) and sometimes *sulfur* (S). Proteins are much more diverse chemically and functionally than lipids and carbohydrates. They can be constructed of any combination of about 20 small building blocks called **amino acids** (Fig. 2.18A).

Amino acids get their name from the fact that they are always formed of an *amine* (or amino) group (HNH, or NH_2) and an acidic *carboxyl* group (COOH). These two functional groups are joined by a carbon atom. Amino acids differ from one another according to the type of molecule (called the *side chain* or *R group*) attached to the central carbon atom. Side chains can be hydrophilic or hydrophobic, they can be large or small, and they can be positively or negatively charged.

To form proteins, amino acids are joined into chains by dehydration synthesis (Fig. 2.18B). The bond between the carbon and nitrogen of adjacent amino acids is a *peptide bond*. A **peptide** is two or more amino acids linked by peptide bonds. The side chains project out from the amino acid chain. Chains shorter than about 100 amino acid molecules are called *polypeptides* (Fig. 2.18C). Chains of amino acids longer than about 100 molecules are called **proteins.**

Of all organic molecules, proteins have the most diverse roles because they can be made with so many different combinations of amino acids. Think of the 20 amino acids as an alphabet for building proteins. One sequence could be SOAP and another SOUP; a small change, just one letter, but a huge difference in function and taste. By varying the "spelling" of amino acids in

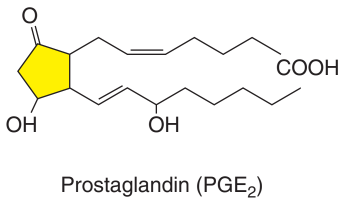
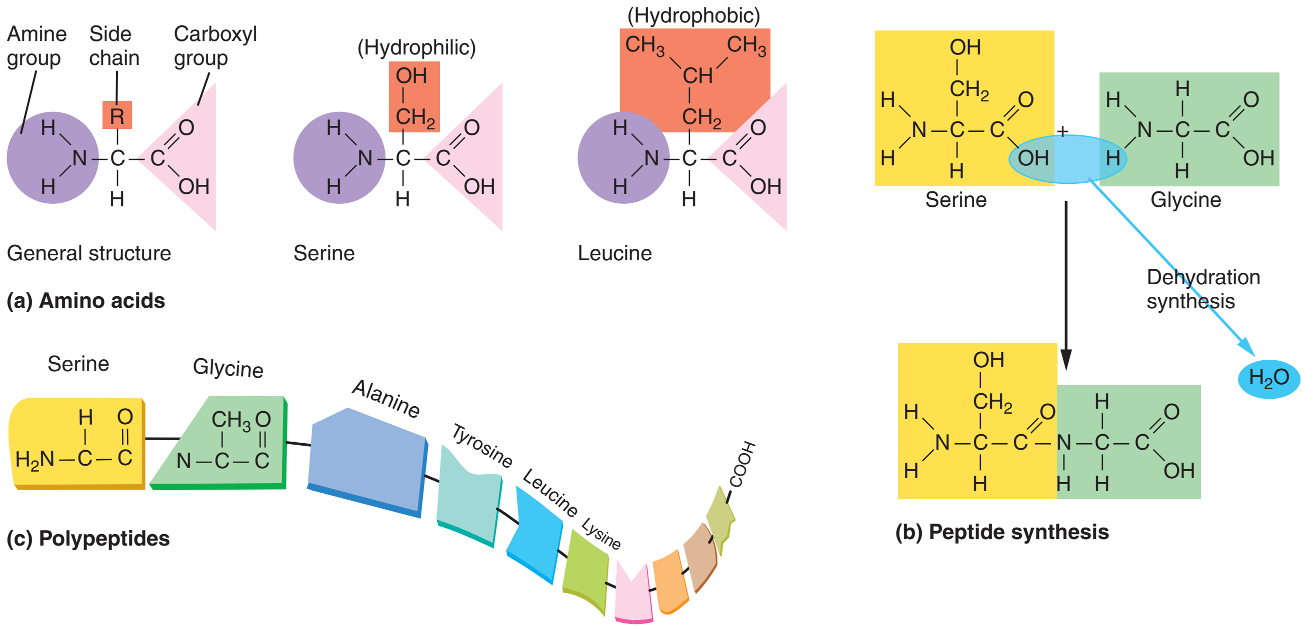


Figure 2.17. Prostaglandins. Prostaglandin E2 (PGE2) is derived from a fatty acid. *Which functional group is COOH?*



(a) Amino acids

(b) Peptide synthesis

(c) Polypeptides

Figure 2.18. Proteins. A. Amino acids contain three functional groups: an amine group, a carboxyl group, and a variable side chain. **B.** Amino acids are monomers joined together into polymers by peptide bonds. **C.** A short polymer of amino acids is called a polypeptide; a longer polymer is called a protein. *How many amine groups are present in the assembled peptide in part B?*

proteins, the body can construct thousands of different proteins using the same amino acid ingredients.

The function of our body proteins is determined in part by their shape. The precise sequence of amino acids in a protein determines how it folds and twists into its characteristic three-dimensional shape, because the side chain of each amino acid will form certain types of bonds only with the side chains of certain other amino acids. The sulfur found in the amino acid *cysteine* (Fig. 2.18B), for instance, forms covalent bonds with other cysteine side chains. Hydrogen and ionic bonds also form between side chains and help fold proteins. This highly specific shaping gives proteins the capacity to perform a wide variety of highly specific tasks.

We can distinguish two main types of proteins according to their shape, structure, and general function. **Fibrous proteins (structural proteins)** are formed of long strands. They are tough, fixed in place, and used to build body structures. The most abundant fibrous protein in the body is *collagen*, which is found in skin, tendons, ligaments, and bones and is the main ingredient in scar tissue. Three interlocking fibrous proteins form collagen (Fig. 2.19A). On the other hand, **globular (functional) proteins** are more rounded molecules that perform specific tasks. **Hemoglobin,** for instance, is a complex globular protein that transports oxygen in the bloodstream (Fig. 2.19B). It contains four separate protein strands, each coiled together into a rounded shape.

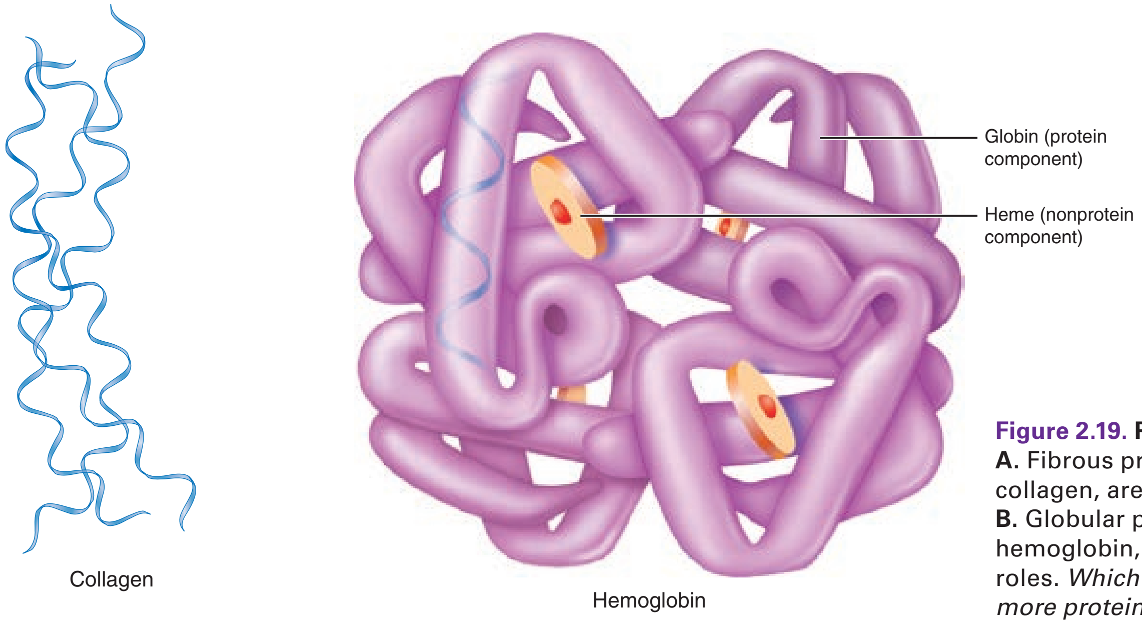
Again, a critical aspect of all proteins is that their *form* is critical to their *function*. Just as the shape of a key must be exactly correct if it is to open a lock, a minor alteration in the shape of a protein can disable it from functioning. For instance, a single amino acid change in the hemoglobin gene can significantly alter the shape of the hemoglobin molecule. This abnormally shaped molecule results in the deformed (crescent- or sickle-shaped) red blood cells that are the result of a genetic disease known as sickle cell anemia.

Case Note

2.23. Notice in the case study figure that 1 cup of lemon juice contains 1 g of protein. What are the building blocks of proteins?

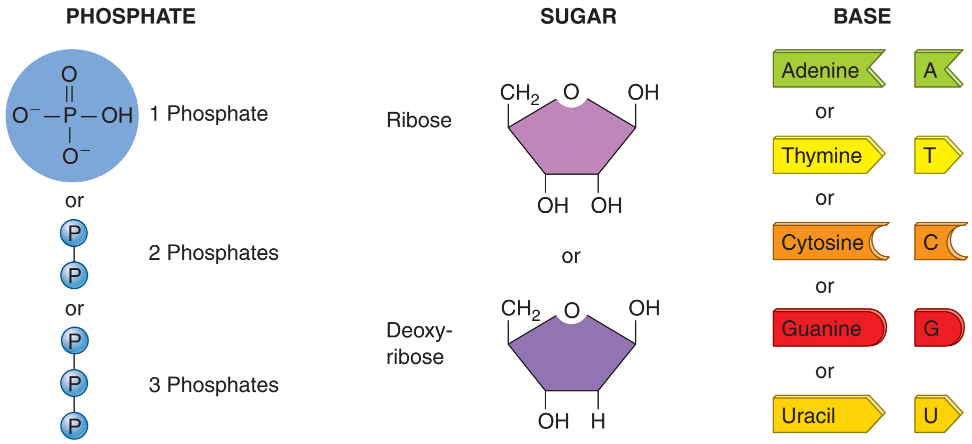
Nucleotides

Nucleotides, the fourth class of organic molecule, play important roles in cell processes such as information storage and the transfer of energy. Each nucleotide contains three building blocks: a simple sugar, a base (a circular molecule containing carbon and nitrogen), and one or more phosphate groups (Fig. 2.20A). The sugar is one of two related sugars, ribose or deoxyribose, that differ only by one hydroxyl group. Five bases

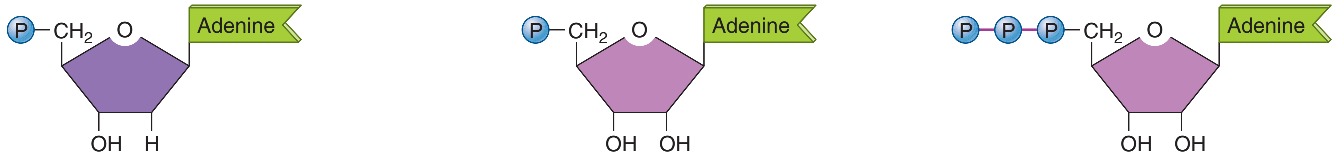


(a) Fibrous protein **(b) Globular protein**

Figure 2.19. Protein shapes. A. Fibrous proteins, such as collagen, are used to build tissue. **B.** Globular proteins, such as hemoglobin, often serve functional roles. *Which protein contains more protein chains: collagen or globin?*



(a) Nucleotide building blocks



Adenine nucleotide (DNA)
= one phosphate + deoxyribose + adenine

Adenine nucleotide (RNA)
= one phosphate + ribose + adenine

ATP (adenosine triphosphate)
= three phosphates + ribose + adenine

(b) Nucleotides

Figure 2.20. Nucleotide structure. A. Nucleotides are built from a phosphate, a sugar, and a base. Two different depictions of each base are shown. **B.** These parts can be assembled into DNA nucleotides, RNA nucleotides, or adenosine triphosphate. *Which nucleotide contains more phosphate groups: the adenine nucleotide or ATP?*

exist: adenine (A), guanine (G), cytosine (C), thymine (T), or uracil (U).

Two types of nucleotides essential to human functioning are nucleic acids and adenosine triphosphate (Fig. 2.20B).

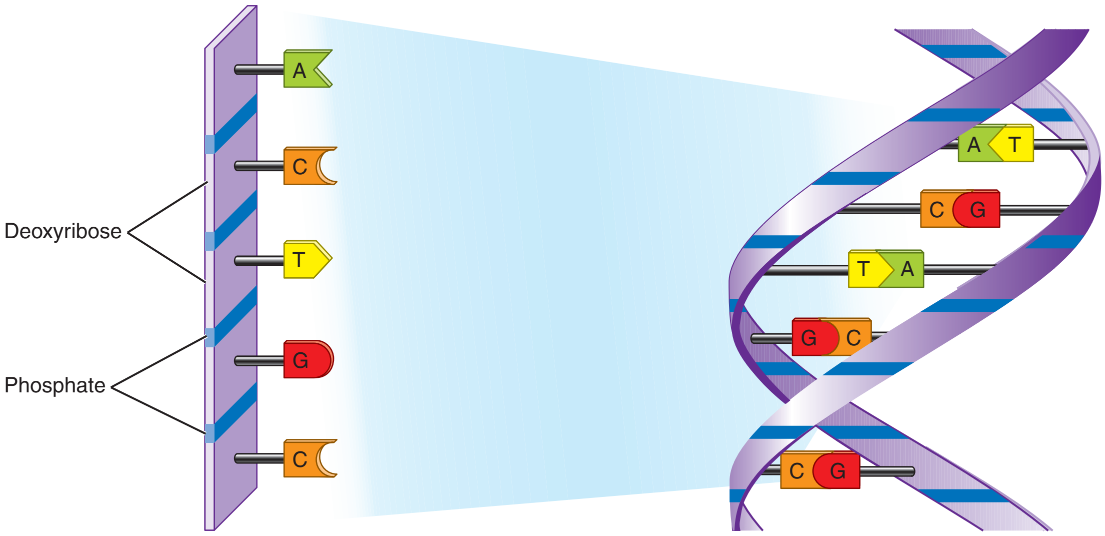
Remember This! **Four classes of organic molecules are: carbohydrates, lipids, proteins, and nucleotides.**

Nucleic Acids Include DNA and RNA

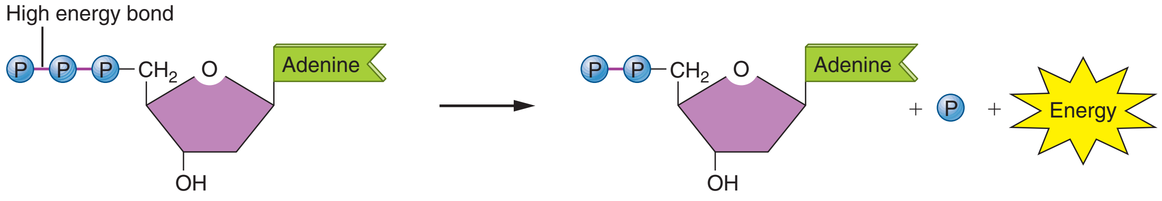
Nucleotides are used to build **nucleic acids,** specifically deoxyribonucleic acid (**DNA**) and ribonucleic acid (**RNA**) (Fig. 2.19B). Notice that nucleic acids contain the three essential building blocks of all nucleotides: either deoxyribose (in deoxyribonucleic acid, DNA) or ribose (in ribonucleic acid, RNA), plus one of the five bases and a single phosphate group. In DNA, the possible bases are adenine (A), guanine (G), thymine (T), and cytosine (C). These are, so to speak, the alphabet of life. In RNA, the possible bases are adenine, guanine, cytosine, and uracil (U).

DNA nucleotides are bound together in long strands, with alternating sugar and phosphate groups forming the DNA backbone and the bases projecting from the strand (Fig. 2.21A). These bases form hydrogen bonds with bases on another strand of DNA to form a "spiral ladder" shape called a double helix. DNA's two strands of nucleotides contain the blueprints to make all of the different body proteins. RNA can act as an agent for DNA, carrying DNA messages, or it can act on its own accord, performing many tasks in the cell.

The sequence of nucleotide molecules in DNA is the genetic code. In the same way that the sequence of letters in this sentence spells out the meaning of the sentence and the sequence of amino acids determines the type of protein, the sequence of nucleotides in DNA and RNA spell out the instructions that our cells follow to build proteins. The sequence is all-important—CTA codes for one amino acid, ACT for a completely different one. Long sequences of these three-letter bases spell out how amino acids are to be strung together to form



(a) DNA encodes genetic information



(b) ATP hydrolysis releases energy

Figure 2.21. Nucleotide function. A. Nucleotides found in DNA or RNA can encode information. **B.** Nucleotides such as ATP store energy in high-energy bonds. *Name the base found in ATP.*

proteins, and proteins, in turn, regulate every aspect of anatomy and physiology. Again, in DNA as in every other aspect of anatomy and physiology, form and function are inseparable.

Adenosine Triphosphate Is Involved in Energy Transfer

The nucleotide **adenosine triphosphate (ATP)** is involved in energy transfer (Fig. 2.20B). ATP is formed of the base *adenine* and the sugar *ribose*, to which are attached three phosphate groups. ATP fulfills a very different purpose than the nucleic acids: ATP is the energy currency of the body (Fig. 2.21B).

Anabolic chemical reactions and many transport mechanisms require energy, and this energy must be in the form of ATP. That is, the energy found in the chemical bonds of nutrients must first be converted into a specific bond in ATP. Most cells use glucose as a nutrient, but some prefer fatty acids or even amino acids. The energy in ATP is contained in a high-energy bonds between two phosphate (PO_4) groups. This bond takes a large amount of energy to create, so it holds a large amount of energy, which is liberated whenever the bond is broken.

Pop Quiz

2.18 If a substance has the chemical formula $C_{18}H_{36}O_2$, is it organic or inorganic?

2.19 Is dehydration synthesis an anabolic or a catabolic reaction?

2.20 How would you classify an amino acid: as a monomer or as a polymer?

2.21 True or false: The assembly of a disaccharide from two monosaccharides generates a water molecule.

2.22 If you dissolve an acidic substance in a solution, will the pH increase or decrease?

2.23 Name the polysaccharide used to store carbohydrate in human muscles.

2.24 List the two most common building blocks of lipids.

2.25 Name a lipid that contains two fatty acid chains.

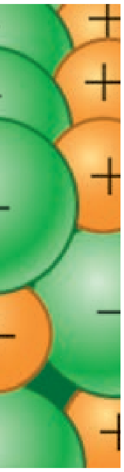
2.26 Name the two functional groups that are identical in every amino acid.

2.27 Name the three building blocks used to construct nucleotides.

Word Parts

Latin/Greek Roots	English Equivalents	Examples
Hydr-/o-	Water	Hydrolysis: breakdown of water
Lip-/o-	Fat	Lipophilic: lipid-loving
-phobia	Fear	Lipophobic: fat-fearing
Phil-	Love	Hydrophilic: water-loving
-lysis	Breakdown	Lipolysis: breakdown of fat
Sacchar-/o-	Sugar	Monosaccharide: one sugar molecule
Mono-	One	Monosaccharide: one sugar molecule
Di-	Two	Disaccharide: sugar molecule containing two saccharide monomers
Tri-	Three	Triglyceride: fat containing three fatty acids and glycerol
Poly-	Many	Polypeptide: protein containing many amino acid subunits
Glyc-/o-	Sugar, glucose	Glycogen: storage form of glucose

Chapter Challenge



CHAPTER RECALL

1. **A positively charged particle is called a(n)**
   a. electron.
   b. proton.
   c. neutron.
   d. isotope.

2. **Two radium atoms contain the same number of electrons and protons, but one atom contains two additional neutrons. These two atoms are**
   a. isotopes.
   b. ions.
   c. polar.
   d. salts.

3. **The hydrogen ion**
   a. has two protons.
   b. is negatively charged.
   c. does not have any electrons.
   d. is an isotope.

4. **Covalent bonds form when**
   a. one atom donates an electron to another atom.
   b. a cation binds with an anion.
   c. polar molecules interact with water.
   d. two atoms share electrons.

5. **Hydrophilic substances are**
   a. frequently dipoles.
   b. often nonpolar.
   c. easily dissolved in fats.
   d. always ions.

6. **Salts are formed by**
   a. two atoms participating in a covalent bond.
   b. two anions.
   c. an anion and a cation.
   d. two cations.

7. **Which of the following substances contains a sugar molecule?**
   a. Triglyceride
   b. Iron
   c. ATP
   d. Carbon

8. **A hydrolysis reaction**
   a. results in the formation of a new water molecule.
   b. assembles monomers into polymers.
   c. liberates energy.
   d. none of the above.

9. **The pH of a solution containing more hydrogen ions than hydroxyl ions might be**
   a. 2.
   b. 7.
   c. 9.
   d. 13.

For each of the following macromolecules, indicate the building blocks that can be used to build them. Each letter may be used more than once.

Macromolecule	**Building Blocks**
10. Protein	a. Glucose
11. Triglyceride	b. Glycerol
12. Glycogen	c. Phosphate
13. Phospholipid	d. Fatty acid
14. ATP	e. Amino acid
15. Starch	f. Adenine (base)
16. Ribonucleic acid	g. Ribose

CONCEPTUAL UNDERSTANDING

Use the periodic table in Figure 2.2 to answer questions 17 through 20.

17. **The atomic mass of phosphorus (P) is 31. How many neutrons are found in its nucleus?**

18. **A common ion in the body is K^+. How many electrons orbit the nucleus of this ion?**

19. **How many electrons will be found in the outer shell of lithium (Li)?**

20. **Name two macromolecules that contain a phosphate group and describe the purpose of the phosphate group in each macromolecule.**

21. Discuss how function is determined by form at the molecular level, using fibrous and functional proteins as examples.

22. Compare and contrast the form and function of phospholipids and triglycerides.

APPLICATION

23. Consider the following reaction:

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

It is possible to reduce the concentration of CO_2 in the blood by hyperventilating (breathing rapidly). Will this action increase or decrease the pH of blood?

24. Draw the structure of a peptide formed with the sequence leucine–serine.

25. Calcium phosphate ($CaPO_4$) is formed by an ionic bond between a calcium cation and a phosphate (PO_4) anion. Phosphate is formed by covalent bonds between one phosphate atom and four oxygen atoms. State which of the following terms describe phosphate, defining each term in your answer: ion, molecule, compound.

You can find the answers to these questions on the student Web site at **http://thepoint.lww.com/McConnellandHull**