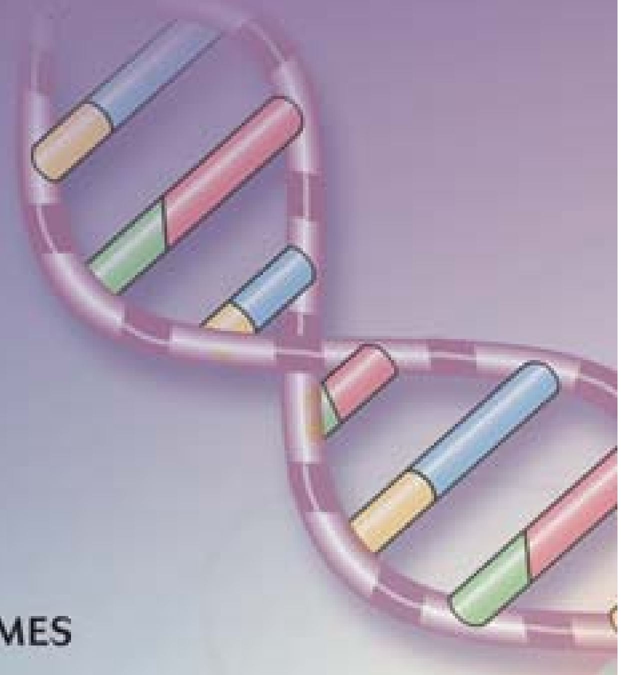
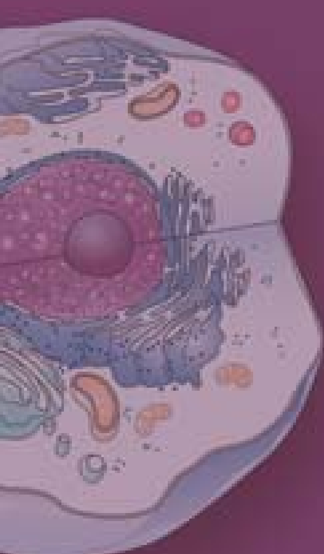




Selected Key Terms

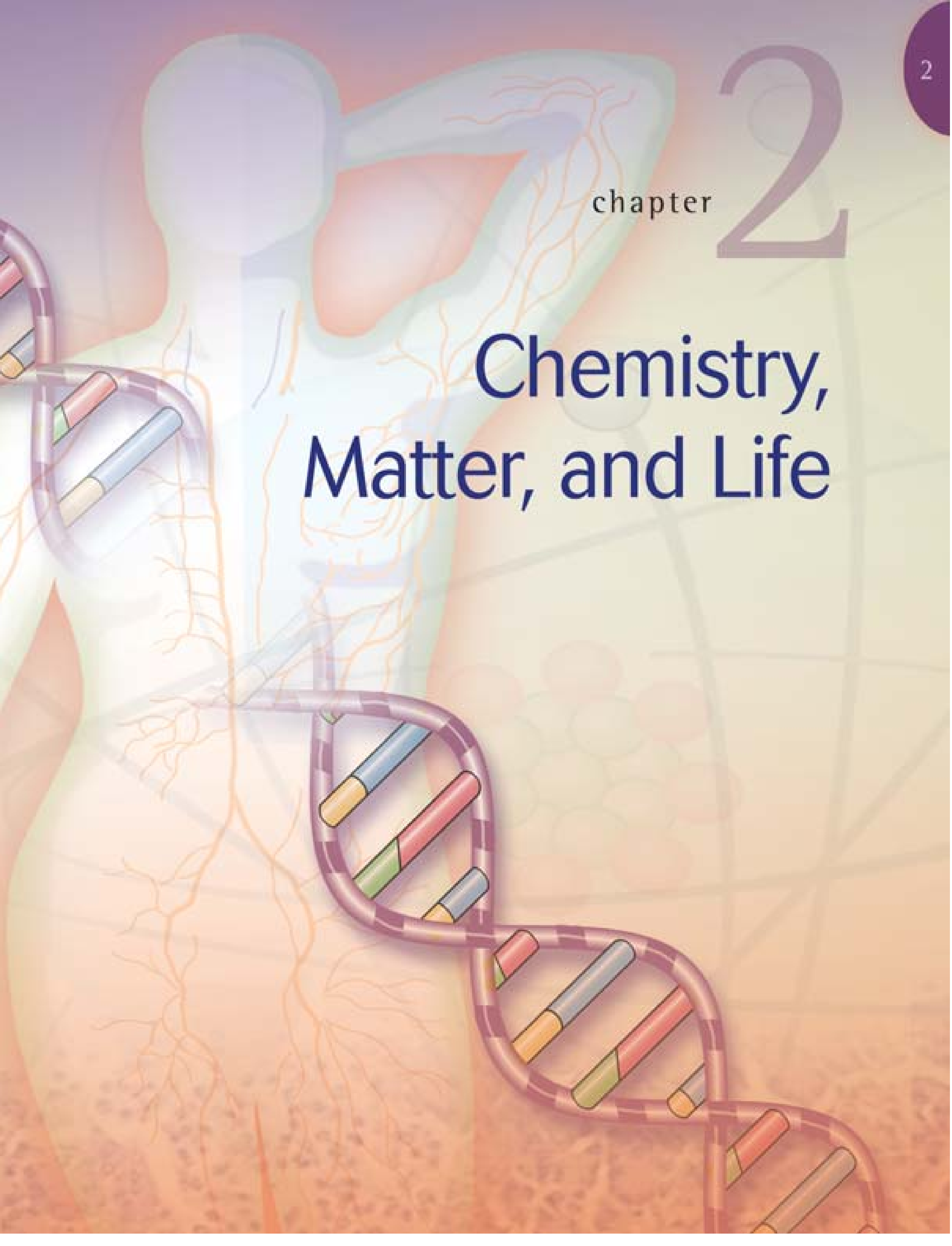
The following terms and other boldface terms in the chapter are defined in the Glossary

acid
alkali
atom
base
buffer
carbohydrate
catalyst
chemistry
colloid
compound
denaturation
electrolyte
electron
element
enzyme
fat
ion
isotope
lipid
molecule
neutron
pH
protein
proton
radioactive
salt
solution
suspension
valence

Learning Outcomes

After careful study of this chapter, you should be able to:

1. Define an element
2. Describe the structure of an atom
3. Differentiate between molecules and compounds
4. Explain why water is so important to the body
5. Define *mixture*; list the three types of mixtures and give two examples of each
6. Differentiate between ionic and covalent bonds
7. Define an electrolyte
8. Define the terms *acid*, *base*, and *salt*
9. Explain how the numbers on the pH scale relate to acidity and basicity (alkalinity)
10. Define *buffer* and explain why buffers are important in the body
11. Define *radioactivity* and cite several examples of how radioactive substances are used in medicine
12. List three characteristics of organic compounds
13. Name the three main types of organic compounds and the building blocks of each
14. Define *enzyme*; describe how enzymes work
15. Show how word parts are used to build words related to chemistry, matter, and life (see Word Anatomy at the end of the chapter)



chapter 2

Chemistry, Matter, and Life

Greater understanding of living organisms has come to us through **chemistry**, the science that deals with the composition and properties of matter. Knowledge of chemistry and chemical changes helps us understand the normal and abnormal functioning of the body. The digestion of food in the intestinal tract, the production of urine by the kidneys, the regulation of breathing, and all other body activities involve the principles of chemistry. The many drugs used to treat diseases are chemicals. Chemistry is used for the development of drugs and for an understanding of their actions in the body.

To provide some insights into the importance of chemistry in the life sciences, this chapter briefly describes elements, atoms and molecules, compounds, and mixtures, which are fundamental forms of matter.

Elements

Matter is anything that takes up space, that is, the materials from which all of the universe is made. Elements are the substances that make up all matter. The food we eat, the atmosphere, water—everything around us, everything we can see and touch, is made of elements. There are 92 naturally occurring elements. (Twenty additional elements have been created in the laboratory.) Examples of elements include various gases, such as hydrogen, oxygen, and nitrogen; liquids, such as mercury used in barometers and other scientific instruments; and many solids, such as iron, aluminum, gold, silver, and zinc. Graphite (the so-called "lead" in a pencil), coal, charcoal, and diamonds are different forms of the element carbon.

Elements can be identified by their names or their chemical symbols, which are abbreviations of the modern or Latin names of the elements. Each element is also identified by its own number, which is based on the structure of its subunits, or atoms. The periodic table is a chart used by chemists to organize and describe the elements. Appendix 3 shows the periodic table and gives some information about how it is used. Table 2-1 lists some elements found in the human body along with their functions.

Table 2•1 Some Common Chemical Elements*

NAME	SYMBOL	FUNCTION
Oxygen	O	Part of water; needed to metabolize nutrients for energy
Carbon	C	Basis of all organic compounds; in carbon dioxide, the waste gas of metabolism
Hydrogen	H	Part of water; participates in energy metabolism, acid–base balance
Nitrogen	N	Present in all proteins, ATP (the energy compound), and nucleic acids (DNA and RNA)
Calcium	Ca	Builds bones and teeth; needed for muscle contraction, nerve impulse conduction, and blood clotting
Phosphorus	P	Active ingredient in the energy-storing compound ATP; builds bones and teeth; in cell membranes and nucleic acids
Potassium	K	Nerve impulse conduction; muscle contraction; water balance and acid–base balance
Sulfur	S	Part of many proteins
Sodium	Na	Active in water balance, nerve impulse conduction, and muscle contraction
Chlorine	Cl	Active in water balance and acid–base balance; found in stomach acid
Iron	Fe	Part of hemoglobin, the compound that carries oxygen in red blood cells

**The elements are listed in decreasing order by weight in the body.*

Atoms

The subunits of elements are **atoms**. These are the smallest complete units of matter. They cannot be broken down or changed into another form by ordinary chemical and physical means. These subunits are so small that millions of them could fit on the sharpened end of a pencil.

Atomic Structure Despite the fact that the atom is such a tiny particle, it has been carefully studied and has been found to have a definite structure. At the center of the atom is a nucleus, which contains positively charged electrical particles called **protons** (PRO-tonz) and noncharged particles called **neutrons** (NU-tronz). Together, the protons and neutrons contribute nearly all of the atom's weight.

In orbit outside the nucleus are **electrons** (e-LEK-tronz) (Fig. 2-1). These nearly weightless particles are negatively charged. It is the electrons that determine how the atom will react chemically. The protons and electrons of an atom always are equal in number, so that the atom as a whole is electrically neutral.

The **atomic number** of an element is equal to the number of protons that are present in the nucleus of each of its atoms. Because the number of protons is equal to the number of electrons, the atomic number also represents the number of electrons whirling around the nucleus. Each element has a specific atomic number. No two elements share the same number. In the Periodic Table of the Elements (see Appendix 3) the atomic number is located at the top of the box for each element.

The positively charged protons keep the negatively charged electrons in orbit around the nucleus by means of the opposite charges on the particles. Positively (+) charged protons attract negatively (−) charged electrons.

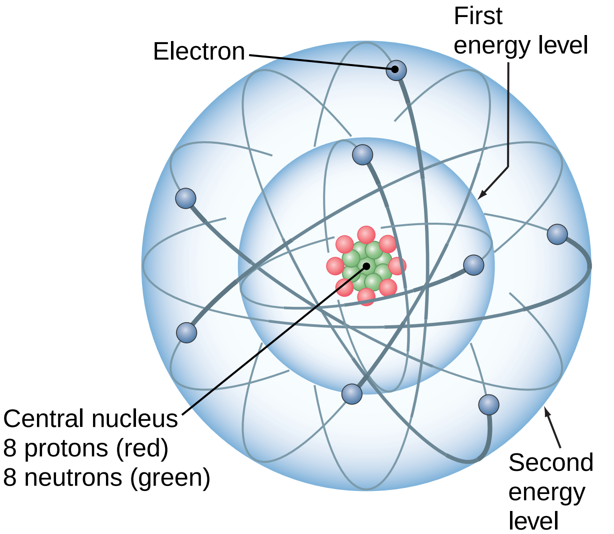


Figure 2-1 **Representation of the oxygen atom.** Eight protons and eight neutrons are tightly bound in the central nucleus. The eight electrons are in orbit around the nucleus, two at the first energy level and six at the second. *ZOOMING IN ◆ How does the number of protons in this atom compare with the number of electrons?*

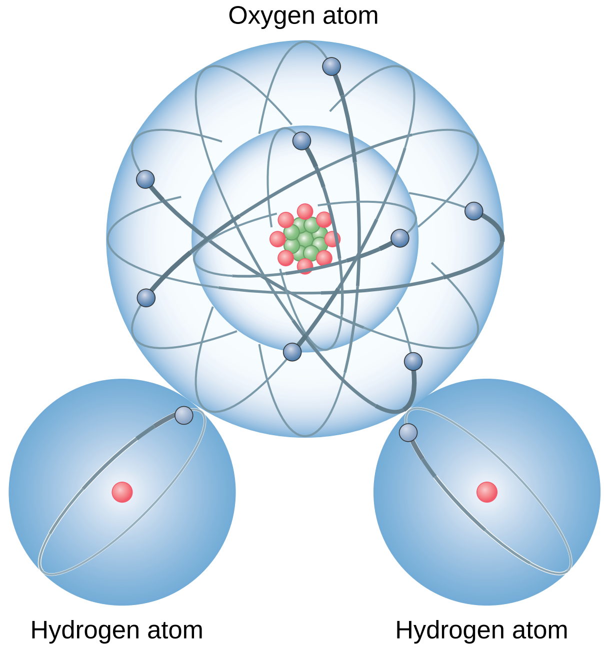


Figure 2-2 **Formation of water.** When oxygen reacts, two electrons are needed to complete the outermost energy level, as shown in this reaction with hydrogen to form water. *ZOOMING IN ◆ How many hydrogen atoms bond with an oxygen atom to form water?*

Checkpoint 2-1 What are atoms?

Checkpoint 2-2 What are three types of particles found in atoms?

Energy Levels The electrons of an atom orbit at specific distances from the nucleus in regions called energy levels. The first energy level, the one closest to the nucleus, can hold only two electrons. The second energy level, the next in distance away from the nucleus, can hold eight electrons.

More distant energy levels can hold more than eight electrons, but they are stable (nonreactive) when they have eight.

The electrons in the energy level farthest away from the nucleus give the atom its chemical characteristics. If the outermost energy level has more than four electrons but less than its capacity of eight, the atom normally completes this level by gaining electrons. In the process, it becomes negatively charged, because it has more electrons than protons. The oxygen atom illustrated in Figure 2-1 has six electrons in its second, or outermost, level. When oxygen enters into chemical reactions, it gains two electrons, as when it reacts with hydrogen to form water (Fig. 2-2). The oxygen atom then has two more electrons than protons.

If the outermost energy level has fewer than four electrons, the atom normally loses those electrons. In so doing, it becomes positively charged, because it now has more protons than electrons.

The number of electrons lost or gained by atoms of an element in chemical reactions is known as the **valence** of that element (from a Latin word that means "strength"). The outermost energy level, which determines the combining properties of the element, is the valence level. Valence is reported as a number with a + or – to indicate whether electrons are lost or gained in chemical reactions. Remember that electrons carry a negative charge, so when an atom gains electrons it becomes negatively charged and when an atom loses electrons it becomes positively charged. For example, the valence of oxygen, which gains two electrons in chemical reactions, is shown as O^{2-}.

Molecules and Compounds

A **molecule** (MOL-eh-kule) is formed when two or more atoms unite on the basis of their electron structures. A molecule can be made of like atoms—the oxygen molecule is made of two identical atoms—but more often a molecule is made of atoms of two or more different elements. For example, a molecule of water (H_2O) contains one atom of oxygen (O) and two atoms of hydrogen (H) (see Fig. 2-2).

Substances composed of two or more different elements are called **compounds.** Molecules are the smallest subunits of a compound. Each molecule of a compound contains the elements that make up that compound in the proper ratio. Some compounds are made of a few elements in a simple combination. For example, the gas carbon monoxide (CO) contains 1 atom of carbon (C) and 1 atom of oxygen (O). Other compounds are very large and complex. Such complexity characterizes many of the compounds found in living organisms. Some proteins, for example, have thousands of atoms.

It is interesting to observe how different a compound is from any of its constituents. For example, a molecule of liquid water is formed from oxygen and hydrogen, both of which are gases. Another example is a crystal sugar, glucose ($C_6H_{12}O_6$). Its constituents include 12 atoms of the gas hydrogen, 6 atoms of the gas oxygen, and 6 atoms of the solid element carbon. The component gases and the solid carbon do not in any way resemble the glucose.

Checkpoint 2-3 What are molecules?

The Importance of Water

Water is the most abundant compound in the body. No plant or animal, including the human, can live very long without it. Water is of critical importance in all physiological processes in body tissues. A deficiency of water, or dehydration (de-hi-DRA-shun), can be a serious threat to health. Water carries substances to and from the cells and makes possible the essential processes of absorption, exchange, secretion, and excretion. What are some of the properties of water that make it such an ideal medium for living cells?

- Water can dissolve many different substances in large amounts. For this reason, it is called the **universal solvent.** Many of the materials needed by the body, such as gases, minerals, and nutrients, dissolve in water to be carried from place to place. Substances, such as salts, that mix with or dissolve in water are described as *hydrophilic* ("water-loving"); those, such as fats, that repel and do not dissolve in water are described as *hydrophobic* ("water-fearing").
- Water is stable as a liquid at ordinary temperatures. Water does not freeze until the temperature drops to 0° C (32° F) and does not boil until the temperature reaches 100° C (212° F). This stability provides a constant environment for body cells. Water can also be used to distribute heat throughout the body and to cool the body by evaporation of sweat from the body surface.
- Water participates in chemical reactions in the body. It is needed directly in the process of digestion and in many of the metabolic reactions that occur in the cells.

Checkpoint 2-4 What is the most abundant compound in the body?

Mixtures: Solutions and Suspensions

Not all elements or compounds combine chemically when brought together. The air we breathe every day is a mixture of gases, largely nitrogen, oxygen, and carbon dioxide, along with smaller percentages of other substances. The constituents in the air maintain their identity, although the proportions of each may vary. Blood plasma is also a mixture in which the various components maintain their identity. The many valuable compounds in the plasma remain separate entities with their own properties. Such combinations are called **mixtures—**blends of two or more substances (Table 2-2).

A mixture formed when one substance dissolves in another is called a **solution.** One example is salt water. In a solution, the component substances cannot be distinguished from each other and they remain evenly distributed throughout; that is, the mixture is homogeneous (ho-mo-JE-ne-us). The dissolving substance, which in the body is water, is the **solvent.** The substance dissolved, salt in the case of salt water, is the **solute.** An **aqueous** (A-kwe-us) **solution** is one in which water is the solvent. Aqueous solutions of glucose, salts, or both of these together are used for intravenous fluid treatments.

In some mixtures, the substance distributed in the background material is not dissolved and will settle out unless the mixture is constantly shaken. This type of non-uniform, or heterogeneous (het-er-o-JE-ne-us), mixture is called a **suspension.** The particles in a suspension are separate from the material in which they are dispersed, and they settle out because they are large and heavy. Examples of suspensions are milk of magnesia, finger paints, and, in the body, red blood cells suspended in blood plasma.

Table 2•2 Mixtures

TYPE	DEFINITION	EXAMPLE
Solution	Homogeneous mixture formed when one substance (solute) dissolves in another (solvent)	Table salt (NaCl) dissolved in water; table sugar (sucrose) dissolved in water
Suspension	Heterogeneous mixture in which one substance is dispersed in another, but will settle out unless constantly mixed	Red blood cells in blood plasma; milk of magnesia
Colloid	Heterogeneous mixture in which the suspended material remains evenly distributed based on the small size and opposing charges of the particles	Blood plasma; cytosol

One other type of mixture is of importance in the body. Some organic compounds form **colloids,** in which the molecules do not dissolve yet remain evenly distributed in the suspending material. The particles have electrical charges that repel each other, and the molecules are small enough to stay in suspension. The fluid that fills the cells (cytosol) is a colloidal suspension, as is blood plasma.

Many mixtures are complex, with properties of solutions, suspensions, and colloidal suspensions. Blood plasma has dissolved compounds, making it a solution. The red blood cells and other formed elements give blood the property of a suspension. The proteins in the plasma give it the property of a colloidal suspension. Chocolate milk also has all three properties.

Checkpoint 2-5 Both solutions and suspensions are types of mixtures. What is the difference between them?

Chemical Bonds

When discussing the structure of the atom, we mentioned the positively charged (+) protons that are located in the nucleus and the equal number of orbiting negatively charged (−) electrons that neutralize the protons (Fig. 2-3 A). Atoms interact, however, to reach a stable number of electrons in the outermost energy level. These chemical interactions alter the neutrality of the atoms and also form a bond between them. In chemical reactions, electrons may be transferred from one atom to another or may be shared between atoms.

Ionic Bonds

When electrons are transferred from one atom to another, the type of bond formed is called an **ionic** (i-ON-ik) **bond.** The sodium atom, for example, tends to lose the single electron in its outermost shell (Fig. 2-3 B), leaving an outermost shell with a stable number of electrons (8). Removal of a single electron from the sodium atom leaves one more proton than electrons, and the atom then has a single net positive charge. The sodium atom in this form is symbolized as Na^+. An atom or group of atoms with a positive or negative charge is called an **ion** (I-on). Any ion that is positively charged is a **cation** (CAT-i-on).

Alternately, atoms can gain electrons so that there are more electrons than protons. Chlorine, which has seven electrons in its outermost energy level, tends to gain one electron to fill the level to its capacity. Such an atom of chlorine is negatively charged (Cl^-) (see Fig. 2-3 B). (Chemists refer to this charged form of chlorine as *chloride.*) Any negatively charged ion is an **anion** (AN-i-on).

Let us imagine a sodium atom coming in contact with a chlorine atom. The chlorine atom gains an electron from the sodium atom, forming an ionic bond. The two newly formed ions (Na^+ and Cl^-), because of their opposite charges, attract each other to produce the compound sodium chloride, ordinary table salt (Fig. 2-3 C).

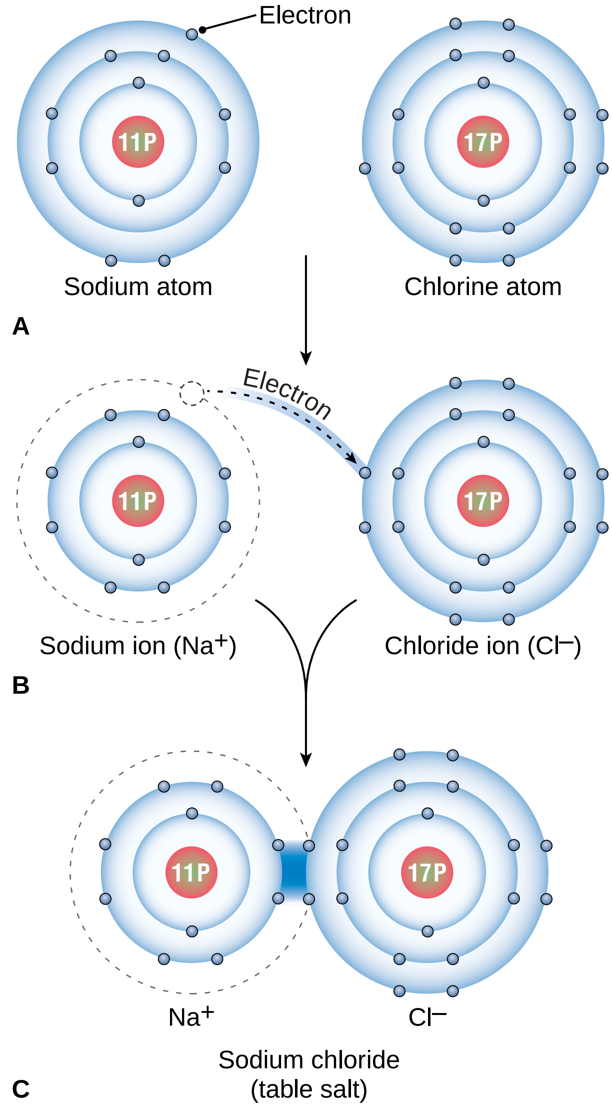


Figure 2-3 Ionic bonding. (A) A sodium atom has 11 protons and 11 electrons. A chlorine atom has 17 protons and 17 electrons. **(B)** A sodium atom gives up one electron to a chlorine atom in forming an ionic bond. The sodium atom now has 11 protons and 10 electrons, resulting in a positive charge of one. The chlorine becomes negatively charged by one, with 17 protons and 18 electrons. **(C)** The sodium ion (Na+) is attracted to the chloride ion (Cl-) in forming the compound sodium chloride (table salt).

Electrolytes Ionically bonded substances, when they go into solution, separate into charged particles. Compounds formed by ionic bonds that release ions when they are in solution are called **electrolytes** (e-LEK-tro-lites). Note that in practice, the term *electrolytes* is also used to refer to the ions themselves in body fluids. Electrolytes include a variety of salts, such as sodium chloride and potassium chloride. They also include acids and bases, which are responsible for the acidity or alkalinity of body fluids, as described later in this chapter. Electrolytes must be present in exactly the right quantities in the fluid within the cell (intracellular fluid) and the fluid outside the cell (extracellular fluid), or very damaging effects will result, preventing the cells in the body from functioning properly.

Ions in the Body Many different ions are found in body fluids. Calcium ions (Ca^{2+}) are necessary for the clotting of blood, the contraction of muscle, and the health of bone tissue. Bicarbonate ions (HCO_3^-) are required for the regulation of acidity and alkalinity of body fluids. The stable condition of the normal organism, homeostasis, is influenced by ions.

Because ions are charged particles, electrolyte solutions can conduct an electric current. Records of electric currents in tissues are valuable indications of the functioning or malfunctioning of tissues and organs. The **electrocardiogram** (e-lek-tro-KAR-de-o-gram) and the **electroencephalogram** (e-lek-tro-en-SEF-ah-lo-gram) are graphic tracings of the electric currents generated by the heart muscle and the brain, respectively (see Chapters 10 and 14).

Checkpoint 2-6 What happens when an electrolyte goes into solution?

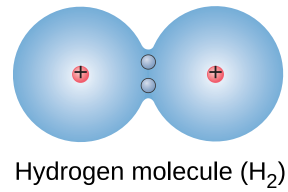


Figure 2-4 **A nonpolar covalent bond.** The electrons involved in the bonding of a hydrogen molecule are equally shared between the two atoms of hydrogen. The electrons orbit evenly around the two. *ZOOMING IN ◆ How many electrons are needed to complete the energy level of each hydrogen atom?*

Covalent Bonds

Although ionic bonds form many chemical compounds, a much larger number of compounds are formed by another type of chemical bond. This bond involves not the exchange of electrons but a sharing of electrons between the atoms in the molecule and is called a **covalent bond.** This name comes from the prefix *co-*, meaning "together," and *valence*, referring to the electrons involved in chemical reactions between atoms. In a covalently bonded molecule, the valence electrons orbit around both of the atoms, making both of them stable. Covalent bonds may involve the sharing of one, two, or three pairs of electrons between atoms.

In some covalently bonded molecules, the electrons are equally shared, as in the case of a hydrogen molecule (H_2) and other molecules composed of atoms of the same element (Fig. 2-4). Electrons may also be shared equally in some

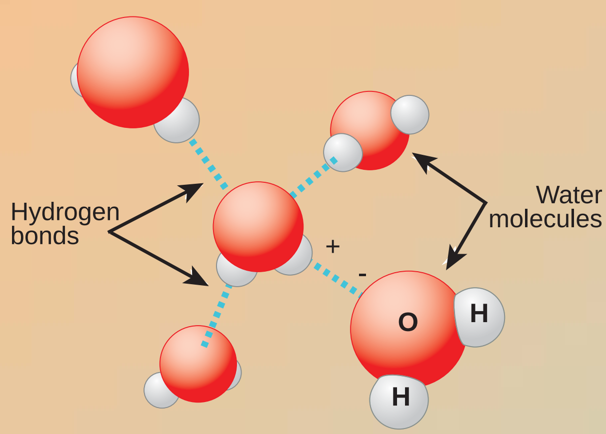
Box 2-1 **A Closer Look**

Hydrogen Bonds: Strength in Numbers

In contrast to ionic and covalent bonds, which hold atoms together, hydrogen bonds hold molecules together. Hydrogen bonds are much weaker than ionic or covalent bonds—in fact, they are more like "attractions" between molecules. While ionic and covalent bonds rely on electron transfer or sharing, hydrogen bonds form bridges between two molecules. A hydrogen bond forms when a slightly positive hydrogen atom in one molecule is attracted to a slightly negative atom in another molecule. Even though a single hydrogen bond is weak, many hydrogen bonds between two molecules can be strong.

Hydrogen bonds hold water molecules together, with the slightly positive hydrogen atom in one molecule attracted to a slightly negative oxygen atom in another. Many of water's unique properties come from its ability to form hydrogen bonds. For example, hydrogen bonds keep water liquid over a wide range of temperatures, which provides a constant environment for body cells.

Hydrogen bonds form not only between molecules but also within large molecules. Hydrogen bonds between regions of the same molecule cause it to fold and coil into a specific shape, as in the process that creates the precise three-dimensional structure of proteins. Because a protein's structure determines its function in the body, hydrogen bonds are essential to protein activity.



Hydrogen bonds. The bonds shown here are holding water molecules together.

molecules composed of different atoms, methane (CH_4), for example. If electrons are equally shared in forming a molecule, the electrical charges are evenly distributed around the atoms and the bond is described as a *nonpolar covalent bond*. That is, no part of the molecule is more negative or positive than any other part of the molecule. More commonly, the electrons are held closer to one atom than the other, as in the case of water (H_2O), shown in Figure 2-2. In a water molecule, the shared electrons are actually closer to the oxygen at any one time making that region of the molecule more negative. Such bonds are called *polar covalent bonds,* because one part of the molecule is more negative and one part is more positive at any one time. Anyone studying biological chemistry (biochemistry) is interested in covalent bonding because carbon, the element that is the basis of organic chemistry, forms covalent bonds with a wide variety of different elements. Thus, the compounds that are characteristic of living things are covalently bonded compounds. For a description of another type of bond, see Box 2-1, Hydrogen Bonds: Strength in Numbers.

Checkpoint 2-7 How is a covalent bond formed?

Compounds: Acids, Bases, and Salts

An **acid** is a chemical substance capable of donating a hydrogen ion (H^+) to another substance. A common example is hydrochloric acid, the acid found in stomach juices:

$$\underset{\text{(hydrochloric acid)}}{HCl} \rightarrow \underset{\text{(hydrogen ion)}}{H^+} + \underset{\text{(chloride ion)}}{Cl^-}$$

A **base** is a chemical substance, usually containing a hydroxide ion (OH^-), that can accept a hydrogen ion. A base is also called an **alkali** (AL-kah-li). Sodium hydroxide, which releases hydroxide ion in solution, is an example of a base:

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

A reaction between an acid and a base produces a **salt**, such as sodium chloride:

$$HCl + NaOH \rightarrow NaCl + H_2O$$

The pH Scale

The greater the concentration of hydrogen ions in a solution, the greater is the acidity of that solution. The greater the concentration of hydroxide ion (OH^-), the greater the basicity (alkalinity) of the solution. Based on changes in the balance of ions in solution, as the concentration of hydrogen ions increases, the concentration of hydroxide ions decreases. Conversely, as the concentration of hydroxide ions increases, the concentration of hydrogen ions decreases. Acidity and alkalinity are indicated by **pH** units, which represent the relative concentrations of hydrogen and hydroxide ions in a solution. The pH units are listed on a scale from 0 to 14, with 0 being the most acidic and 14 being the most basic (Fig. 2-5). A pH of 7.0 is neutral. At pH 7.0 the solution has an equal number of hydrogen and hydroxide ions. Pure water has a pH of 7.0. Solutions that measure less than 7.0 are acidic; those that measure above 7.0 are alkaline (basic).

Because the pH scale is based on multiples of 10, each pH unit on the scale represents a 10-fold change in the number of hydrogen and hydroxide ions present. A solution registering 5.0 on the scale has 10 times the number of hydrogen ions as a solution that registers 6.0. The pH 5.0 solution also has one tenth the number of hydroxide ions as the solution of pH 6.0. A solution registering 9.0 has one tenth the number of hydrogen ions and 10 times the number of hydroxide ions as one registering 8.0. Thus, the lower the pH reading, the greater is the acidity, and the higher the pH, the greater is the alkalinity.

Blood and other body fluids are close to neutral but are slightly on the alkaline side, with a pH range of

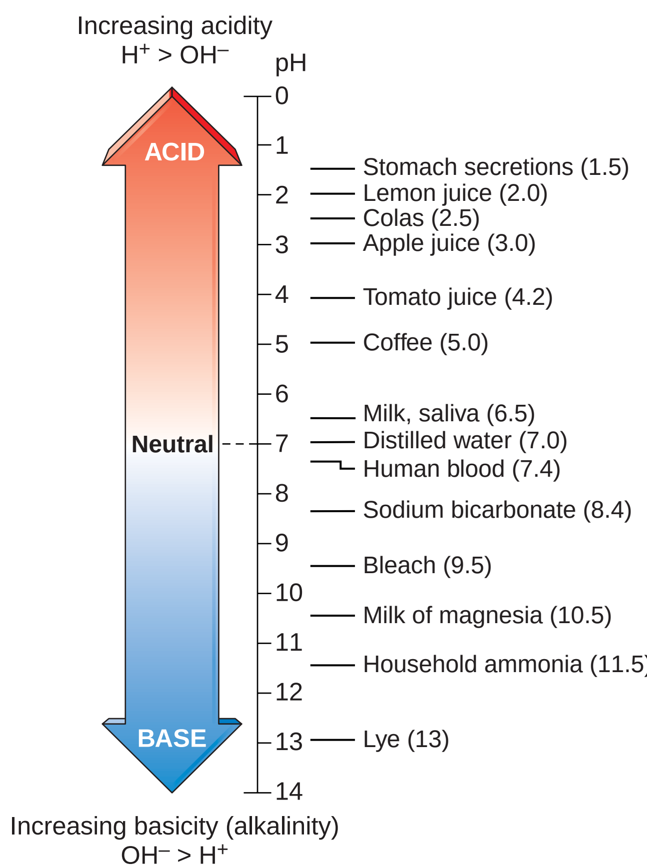


Figure 2-5 **The pH scale.** Degree of acidity or alkalinity is shown in pH units. This scale also shows the pH of some common substances. *ZOOMING IN ✦ What happens to the amount of hydroxide ion (OH^-) present in a solution when the amount of hydrogen ion (H^+) increases?*

7.35 to 7.45. Urine averages pH 6.0 but may range from 4.6 to 8.0 depending on body conditions and diet. Figure 2-5 shows the pH of some other common substances.

Because body fluids are on the alkaline side of neutral, the body may be considered to be in an acidic state even if the pH does not drop below 7.0. For example, if the pH falls below 7.35 but is still greater than 7.0, one is described as being in an acidic state known as *acidosis*. Thus, within a narrow range of the pH scale, physiologic acidity may differ from acidity from a chemical standpoint.

An increase in pH to readings greater than 7.45 is termed *alkalosis*. Any shifts in pH to readings above or below the normal range can be dangerous, even fatal.

Buffers

If a person is to remain healthy, a delicate balance must exist within the narrow limits of acidity and alkalinity of body fluids. This balanced chemical state is maintained in large part by **buffers.** Chemicals that serve as buffers form a system that prevents sharp changes in hydrogen ion concentration and thus maintains a relatively constant pH. Buffers are important in maintaining stability in the pH of body fluids. More information about body fluids, pH, and buffers can be found in Chapter 21.

Checkpoint 2-8 The pH scale is used to measure acidity and alkalinity of fluids. What number is neutral on the pH scale? What kind of compound measures lower than this number? Higher?

Checkpoint 2-9 What is a buffer?

Isotopes and Radioactivity

Elements may exist in several forms, each of which is called an **isotope** (I-so-tope). These forms are alike in their numbers of protons and electrons, but differ in their atomic weights because of differing numbers of neutrons in the nucleus. The most common form of oxygen, for example, has eight protons and eight neutrons in the nucleus, giving the atom an atomic weight of 16 atomic mass units (amu). But there are some isotopes of oxygen with only six or seven neutrons in the nucleus and others with 9-11 neutrons. The isotopes of oxygen thus range in weight from 14 to 19 amu.

Some isotopes are stable and maintain constant characteristics. Others disintegrate (fall apart) and give off rays of atomic particles. Such isotopes are said to be **radioactive.** Radioactive elements may occur naturally, as is the case with isotopes of the very heavy elements radium and uranium. Others may be produced artificially by placing the atoms of lighter, non-radioactive elements in accelerators that smash their nuclei together.

Use of Radioactive Isotopes

The rays given off by some radioactive elements, also called *radioisotopes*, are used in the treatment of cancer because they have the ability to penetrate and destroy tissues. Radiation therapy is often given by means of machines that are able to release tumor-destroying particles. The sensitivity of the younger, dividing cells in a growing cancer allows selective destruction of these abnormal cells with minimal damage to normal tissues. Modern radiation instruments produce tremendous amounts of energy (in the multimillion electron-volt range) and yet can destroy

Box 2-2 **Hot Topics**

Radioactive Tracers: Medicine Goes Nuclear

Like radiography, computed tomography, and MRI, **nuclear medicine imaging** (NMI) offers a noninvasive way to look inside the body. An excellent diagnostic tool, NMI shows not only structural details but also provides information about body function. NMI can diagnose cancer, stroke, and heart disease earlier than techniques that provide only structural information.

NMI uses **radiotracers**, radioactive substances that specific organs absorb. For example, radioactive iodine is used to image the thyroid gland, which absorbs more iodine than any other organ. After a patient ingests, inhales, or is injected with a radiotracer, a device called a gamma camera detects the radiotracer in the organ under study and produces a picture, which is used in making a diagnosis. Radiotracers are broken down and eliminated through urine or feces, so they leave the body quickly. A patient's exposure to radiation in NMI is usually considerably lower than with x-ray or CT scan.

Three NMI techniques are **positron emission tomography** (PET), **bone scanning**, and the **thallium stress test**. PET is often used to evaluate brain activity by measuring the brain's use of radioactive glucose. PET scans can reveal brain tumors because tumor cells are often more metabolically active than normal cells and thus absorb more radiotracer. Bone scanning detects radiation from a radiotracer absorbed by bone tissue with an abnormally high metabolic rate, such as a bone tumor. The thallium stress test is used to diagnose heart disease. A nuclear medicine technologist injects the patient with radioactive thallium, and a gamma camera images the heart during exercise and then rest. When compared, the two sets of images help to evaluate blood flow to the working, or "stressed," heart.

deep-seated cancers without causing serious skin reactions.

In radiation treatment, a radioactive isotope, such as cobalt 60, is sealed in a stainless steel cylinder and mounted on an arm or crane. Beams of radioactivity are then directed through a porthole to the area to be treated. Implants containing radioactive isotopes in the form of needles, seeds, or tubes also are widely used in the treatment of different types of cancer.

In addition to its therapeutic values, irradiation is extensively used in diagnosis. X-rays penetrate tissues and produce an image of their interior on a photographic plate. Radioactive iodine and other "tracers" taken orally or injected into the bloodstream are used to diagnose abnormalities of certain body organs, such as the thyroid gland (see Box 2-2, Radioactive Tracers: Medicine Goes Nuclear). Rigid precautions must be followed by healthcare personnel to protect themselves and the patient when using radiation in diagnosis or therapy because the rays can destroy both healthy and diseased tissues.

Checkpoint 2-10 Some isotopes are stable; others break down to give off atomic particles. What word is used to describe isotopes that give off radiation?

Chemistry of Living Matter

Of the 92 elements that exist in nature, only 26 have been found in living organisms. Most of these are elements that are light in weight. Not all are present in large quantity. Hydrogen, oxygen, carbon, and nitrogen are the elements that make up about 96% of the body by weight (Fig. 2-6). Nine additional elements, calcium, sodium, potassium, phosphorus, sulfur, chlorine, magnesium, iron, and iodine make up most of the remaining 4% of the elements in the body. The remaining 13, including zinc, selenium, copper, cobalt, chromium, and others, are present in extremely small (trace) amounts totaling about 0.1% of body weight.

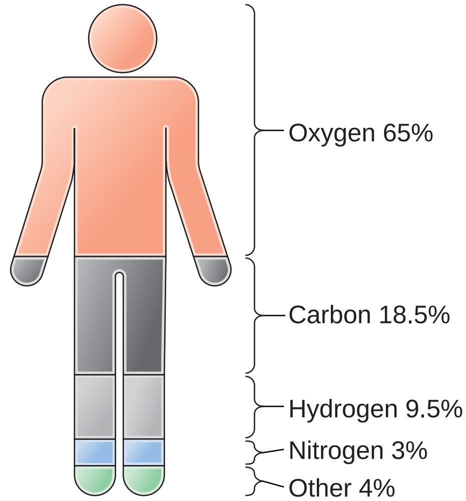


Figure 2-6 **Chemical composition of the body by weight.**

Organic Compounds

The chemical compounds that characterize living things are called **organic compounds.** All of these contain the element **carbon.** Because carbon can combine with a variety of different elements and can even bond to other carbon atoms to form long chains, most organic compounds consist of large, complex molecules. The starch found in potatoes, the fat in the tissue under the skin, hormones, and many drugs are examples of organic compounds. These large molecules are often formed from simpler molecules called *building blocks,* which bond together in long chains.

The main types of organic compounds are carbohydrates, lipids, and proteins. (Another category, the nucleic acids, which are important in cellular functions, are discussed in Chapter 3.) All of these organic compounds contain carbon, hydrogen, and oxygen as their main ingredients.

Carbohydrates, lipids, and proteins, in addition to minerals and vitamins, must be taken in as part of a normal diet. These compounds are discussed further in Chapters 19 and 20.

Checkpoint 2-11 Where are organic compounds found?

Checkpoint 2-12 What element is the basis of organic chemistry?

Carbohydrates The basic units of carbohydrates are simple sugars, or **monosaccharides** (mon-o-SAK-ah-rides) (Fig. 2-7 A). **Glucose** (GLU-kose), a simple sugar that circulates in the blood as a nutrient for cells, is an example of a monosaccharide. Two simple sugars may be linked together to form a **disaccharide** (Fig. 2-7 B), as represented by sucrose, table sugar. More complex carbohydrates, or **polysaccharides** (Fig. 2-7 C), consist of many simple sugars linked together with multiple side chains. Examples of polysaccharides are starch, which is manufactured in plant cells, and **glycogen** (GLI-ko-jen), a storage form of glucose found in liver cells and skeletal muscle cells. Carbohydrates in the form of sugars and starches are important sources of energy in the diet.

Lipids **Lipids** are a class of organic compounds mainly found in the body as **fat.** Fats provide insulation for the body and protection for organs. In addition, fats are the main form in which energy is stored.

Simple fats are made from a substance called **glycerol** (GLIS-er-ol), commonly known as glycerin, in combination with fatty acids (Fig. 2-8 A). One fatty acid is attached

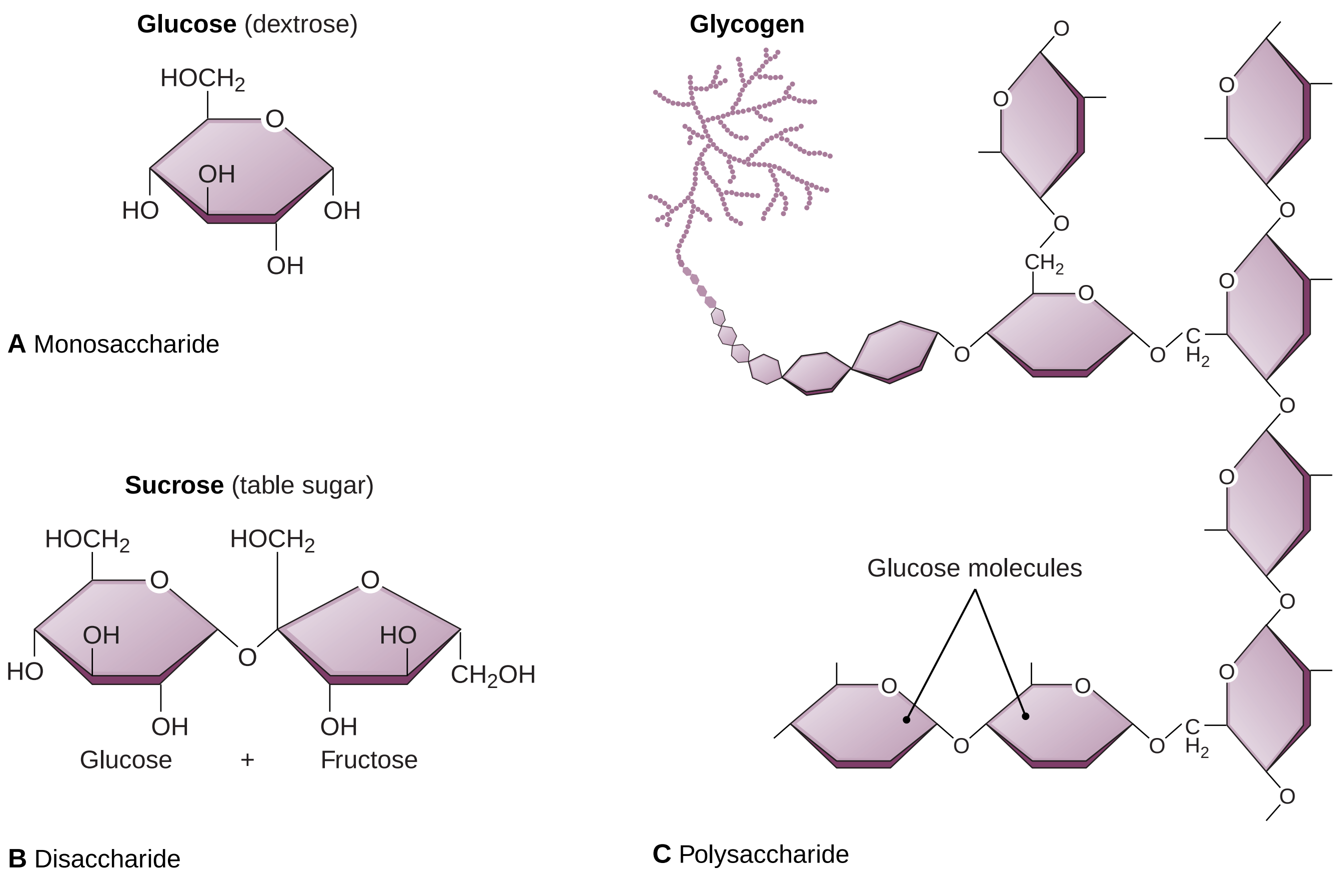


Figure 2-7 **Examples of carbohydrates.** A monosaccharide **(A)** is a simple sugar. A disaccharide **(B)** consists of two simple sugars linked together, whereas a polysaccharide **(C)** consists of many simple sugars linked together in chains. *ZOOMING IN ✦ What are the building blocks of disaccharides and polysaccharides?*

to each of the three carbon atoms in glycerol, so simple fats are described as **triglycerides** (tri-GLIS-er-ides). **Phospholipids** (fos-fo-LIP-ids) are complex lipids containing the element phosphorus. Among other functions, phospholipids make up a major part of the membrane around living cells. **Steroids** are lipids that contain rings of carbon atoms. They include **cholesterol** (ko-LES-ter-ol), another component of cell membranes (Fig. 2-8 B); the steroid hormones, such as cortisol, produced by the adrenal gland; the sex hormones, such as testosterone, produced by the testes; and estrogen and progesterone, produced by the ovaries.

Proteins All **proteins** (PRO-tenes) contain, in addition to carbon, hydrogen, and oxygen, the element **nitrogen** (NI-tro-jen). They may also contain sulfur or phosphorus. Proteins are the structural materials of the body, found in muscle, bone, and connective tissue. They also make up the pigments that give hair, eyes, and skin their color. It is protein that makes each individual physically distinct from others.

Proteins are composed of building blocks called **amino** (ah-ME-no) **acids** (Fig. 2-9 A). Although there are only about 20 different amino acids found in the body, a vast number of proteins can be made by linking them together in different sized molecules and in different combinations.

Each amino acid contains an acid group (COOH) and an amino group (NH_2), the part of the molecule that has the nitrogen. Many amino acids link together to form a polypeptide, which is then arranged into a particular shape. The polypeptide chain is coiled into a helix and may then be pleated or folded back on itself. Several chains also may be folded together (see Fig. 2-9 B). The overall shape of a protein is important to its function, as can be seen in the activity of enzymes.

> **Checkpoint 2-13** What are the three main categories of organic compounds?

Enzymes **Enzymes** (EN-zimes) are proteins that are essential for metabolism. They serve as **catalysts** in the hun-

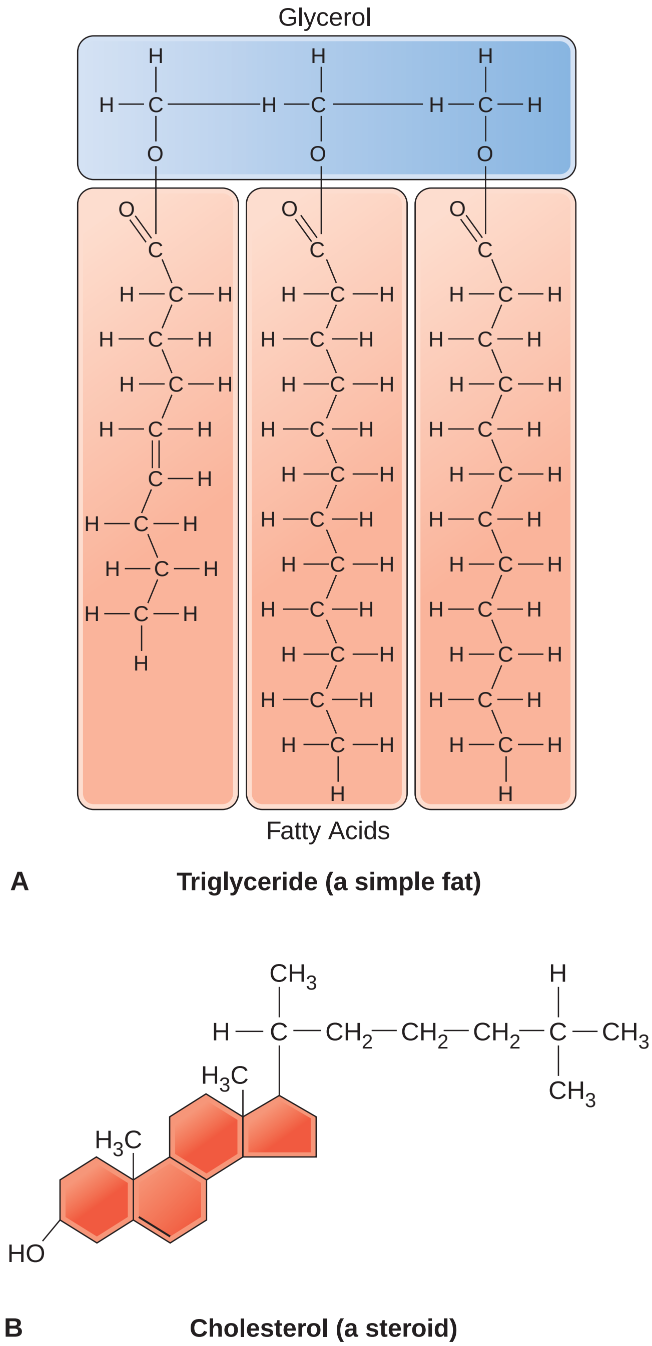


Figure 2-8 Lipids. (A) A triglyceride, a simple fat, contains glycerol combined with three fatty acids. **(B)** Cholesterol is a type of steroid, a lipid that contains rings of carbon atoms. *ZOOMING IN ✦ How many carbon atoms are in glycerol?*

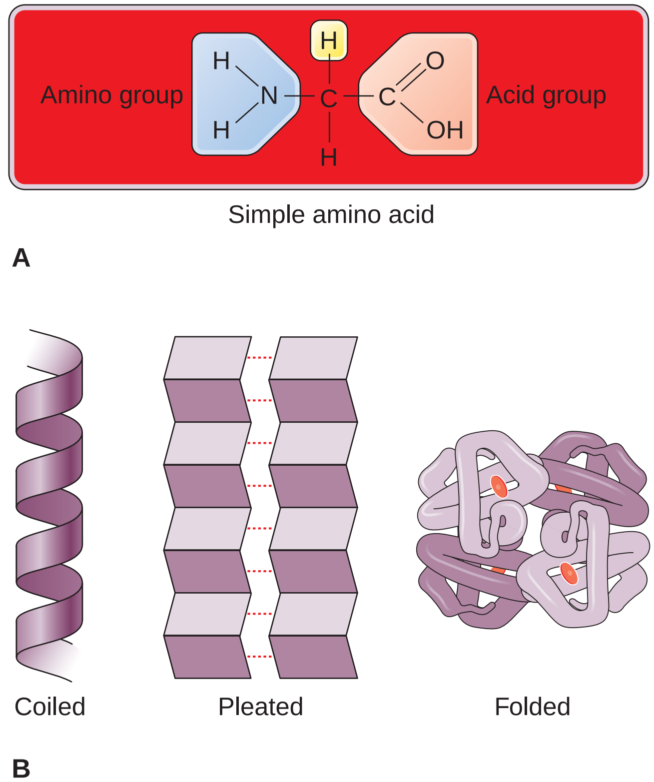


Figure 2-9 Proteins. (A) Amino acids are the building blocks of proteins. **(B)** Some shapes of proteins. *ZOOMING IN ✦ What part of an amino acid contains nitrogen?*

dreds of reactions that take place within cells. Without these catalysts, which speed the rate of chemical reactions, metabolism would not occur at a fast enough rate to sustain life. Because each enzyme works only on a specific substance, or **substrate**, and does only one specific chemical job, many different enzymes are needed. Like all catalysts, enzymes take part in reactions only temporarily; they are not used up or changed by the reaction. Therefore, they are needed in very small amounts. Many of the vitamins and minerals required in the diet are parts of enzymes.

The shape of the enzyme is important in its action. The enzyme's form must match the shape of the substrate or substrates the enzyme combines with in much the same way as a key fits a lock. This so-called "lock-and-key" mechanism is illustrated in Figure 2-10. Harsh conditions, such as extremes of temperature or pH, can alter the shape of an enzyme and stop its action. The alteration of any protein so that it can no longer function is termed **denaturation**. Such an event is always harmful to the cells.

You can usually recognize the names of enzymes because, with few exceptions, they end with the suffix *-ase*. Examples are lipase, protease, and oxidase. The first part of the name usually refers to the substance acted on or the type of reaction in which the enzyme is involved.

Checkpoint 2-14 Enzymes are proteins that act as catalysts. What is a catalyst?

For a description of professions that require knowledge of chemistry, see Box 2-3, Pharmacists and Pharmacy Technicians.

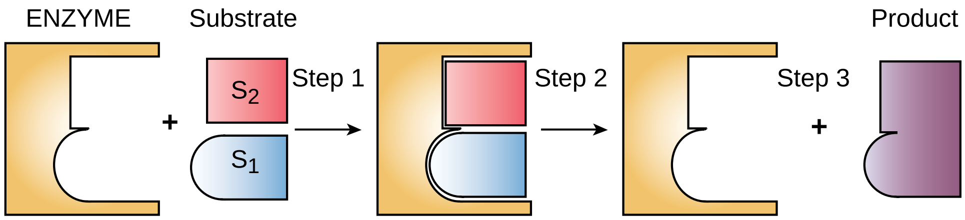


Figure 2-10 **Diagram of enzyme action.** The enzyme combines with substrate 1 (S_1) and substrate 2 (S_2). Once a new product is formed from the substrates, the enzyme is released unchanged. *ZOOMING IN ✦ How does the shape of the enzyme before the reaction compare with its shape after the reaction.*

Box 2-3 • Health Professions

Pharmacists and Pharmacy Technicians

Medications are chemicals designed to treat illness and improve quality of life. The role of pharmacists and pharmacy technicians is to ensure that patients receive the correct medication and the education they need to use it effectively and derive the intended health benefits.

As key members of the healthcare team, pharmacists need a strong clinical background with a thorough understanding of chemistry, anatomy, and physiology. Pharmacists not only dispense prescription medications and monitor patients' responses to them, they also educate patients about their appropriate use. They share their expertise with other health professionals and also participate in clinical research on drugs and their effects.

Pharmacy technicians also require a thorough understanding of chemistry, anatomy, and physiology to assist pharmacists with their duties. State rules and regulations vary, but pharmacy technicians may perform many of the tasks related to dispensing medications, such as preparing them and packaging them with appropriate labels and instructions for use.

Most pharmacists and pharmacy technicians work in retail pharmacies, whereas others work in hospitals and long-term care facilities. Job prospects are promising because of the growing need for healthcare. In fact, pharmacy is projected to be one of the fastest growing careers in the United States. For more information about careers in pharmacy, contact the American Association of Colleges of Pharmacy.

Word Anatomy

Medical terms are built from standardized word parts (prefixes, roots, and suffixes). Learning the meanings of these parts can help you to remember words and interpret unfamiliar terms.

WORD PART	MEANING	EXAMPLE
Molecules and Compounds		
hydr/o	water	*Dehydration* is a deficiency of water.
phil	to like	*Hydrophilic* substances "like" water—they mix with or dissolve in it.
-phobia	fear	*Hydrophobic* substances "fear" water—they repel and do *not* dissolve in it.
hom/o	same	*Homogeneous* mixtures are the same throughout.
heter/o-	different	*Heterogeneous* solutions are different (not uniform) throughout.
aqu/e	water	In an *aqueous* solution, water is the solvent.
Chemical Bonds		
co-	together	*Covalent* bonds form when atoms share electrons.
Chemistry of Living Matter		
sacchar/o	sugar	A *monosaccharide* consists of one simple sugar.
mon/o-	one	In *monosaccharide*, "mono-" refers to one.
di-	twice, double	A *disaccharide* consists of two simple sugars.
poly-	many	A *polysaccharide* consists of many simple sugars.
glyc/o	sugar, glucose, sweet	*Glycogen* is a storage form of glucose. It breaks down to release (generate) glucose.
tri-	three	*Triglycerides* have one fatty acid attached to each of three carbon atoms.

WORD PART	MEANING	EXAMPLE
Chemistry of Living Matter		
de-	remove	*Denaturation* of a protein removes its ability to function (changes its nature).
-ase	suffix used in naming enzymes	A *lipase* is an enzyme that acts on lipids.

Summary

I. Elements—substances from which all matter is made

A. Atoms—subunits of elements
1. Atomic structure
   a. Protons—positively charged particles in the nucleus
   b. Neutrons—noncharged particles in the nucleus
   c. Electrons—negatively charged particles in energy levels around the nucleus
2. Energy levels—orbits that hold electrons at specific distances from the nucleus
   a. Valence—number of electrons lost or gained in chemical reactions

II. Molecules and compounds

1. Molecules—combinations of two or more atoms
2. Compounds—substances composed of different elements

A. The importance of water—solvent; stable; essential for metabolism

B. Mixtures: solutions and suspensions
1. Mixtures: blend of two or more substances
2. Solution: substance (solute) remains evenly distributed in solvent (*e.g.*, salt in water); homogeneous
3. Suspension—material settles out of mixture on standing (*e.g.*, red cells in blood plasma); heterogeneous
4. Colloid—particles do not dissolve but remain suspended (*e.g.*, cytosol)

III. Chemical bonds

A. Ionic bonds—formed by transfer of electrons from one atom to another
1. Electrolytes
   a. Ionically bonded substances
   b. Separate in solution into charged particles (ions); cation positive and anion negative
   c. Conduct electric current
2. Ions in body fluids important for proper function

B. Covalent bonds—formed by sharing of electrons between atoms
1. Nonpolar—equal sharing of electrons (*e.g.*, hydrogen gas, H_2)
2. Polar—unequal sharing of electrons (*e.g.*, water, H_2O)

IV. Compounds: acids, bases and salts

1. Acids—donate hydrogen ions
2. Bases—accept hydrogen ions
3. Salts—formed by reaction between acid and base

A. The pH scale
1. Measure of acidity or alkalinity of a solution
2. Scale goes from 0 to 14
   a. 7 is neutral; below 7 is acidic; above 7 is alkaline (basic)

B. Buffer—maintains constant pH of a solution

V. Isotopes and radioactivity

1. Isotopes—forms of an element that differ in atomic weights (number of neutrons)
   a. Radioactive isotope gives off rays of atomic particles

A. Use of radioactive isotopes
1. Cancer therapy
2. Diagnosis—tracers, x-rays

VI. Chemistry of living matter

A. Organic compounds—all contain carbon
1. Carbohydrates (*e.g.*, sugars, starches); made of simple sugars (monosaccharides)
2. Lipids (*e.g.*, fats, steroids); fats made of glycerol and fatty acids
3. Proteins (*e.g.*, structural materials, enzymes); made of amino acids
   a. Enzymes—organic catalysts

Questions for Study and Review

Building Understanding

Fill in the blanks

1. The basic units of matter are ______.
2. The atomic number is the number of______ in an atom's nucleus.
3. A mixture of solute dissolved in solvent is called a(n) ______.
4. Blood has a pH of 7.35 to 7.45. Gastric juice has a pH of about 2.0. The more alkaline fluid is ______.
5. Proteins that catalyze metabolic reactions are called ______.

Matching

Match each numbered item with the most closely related lettered item.

___6. A simple carbohydrate such as glucose
___7. A complex carbohydrate such as glycogen
___8. An important component of cell membranes
___9. A hormone such as estrogen
___10. The basic building block of protein

a. polysaccharide
b. phospholipid
c. steroid
d. amino acid
e. monosaccharide

Multiple choice

___11. Red blood cells "floating" in plasma are an example of a mixture called a
a. compound
b. suspension
c. colloid
d. solution

___12. The most abundant compound in the body is
a. carbohydrate
b. protein
c. lipid
d. water

___13. A compound that releases ions when it is in solution is called a(n)
a. solvent
b. electrolyte
c. anion
d. colloid

___14. A chemical capable of donating hydrogen ions to other substances is called a(n)
a. acid
b. base
c. salt
d. catalyst

___15. Organic compounds always contain the element
a. oxygen
b. carbon
c. nitrogen
d. phosphorus

Understanding Concepts

16. Compare and contrast the following terms:
    a. element and atom
    b. molecule and compound
    c. proton, neutron, and electron
    d. anion and cation
    e. ionic bond and covalent bond
    f. acid and base
17. What are some of the properties of water that make it an ideal medium for living cells?
18. Explain the importance of ions in the structure and function of the human body.
19. What is pH? Discuss the role of buffers in maintaining pH homeostasis in the body.
20. Compare and contrast carbohydrates and proteins.
21. Describe three different types of lipid.
22. Define the term *enzyme* and discuss the relationship between enzyme structure and enzyme function.

Conceptual Thinking

23. Based on your understanding of strong acids and bases, why does the body have to be kept at a close-to-neutral pH?
24. Mrs. Alvarez has thyroid cancer and is undergoing radiation therapy. During one of her treatments she tells you that she had hoped her initial "thyroid scan" would have killed all of the cancer. Explain the difference between radiation therapy and nuclear medicine imaging.
25. Why do we need enzymes, when usually heat is used to speed up chemical reactions?