CHAPTER 2

The Chemical Basis of Life

Coral reefs are among the most diverse ecosystems on Earth. They are formed from the gradual buildup of the calcium carbonate skeletons of small coral animals. As you can see in the photo below, these structurally diverse habitats provide havens for a huge diversity of fish and other marine organisms. But in recent years, something in the air is threatening coral reefs. How might a chemical compound in the air harm such a vibrant ecosystem? The answer is chemistry. When carbon dioxide dissolves in water, it reacts with water to form an acid, which then makes the water more acidic. Later in the chapter we will see how scientists are exploring the effects of this ocean acidification on coral reefs.

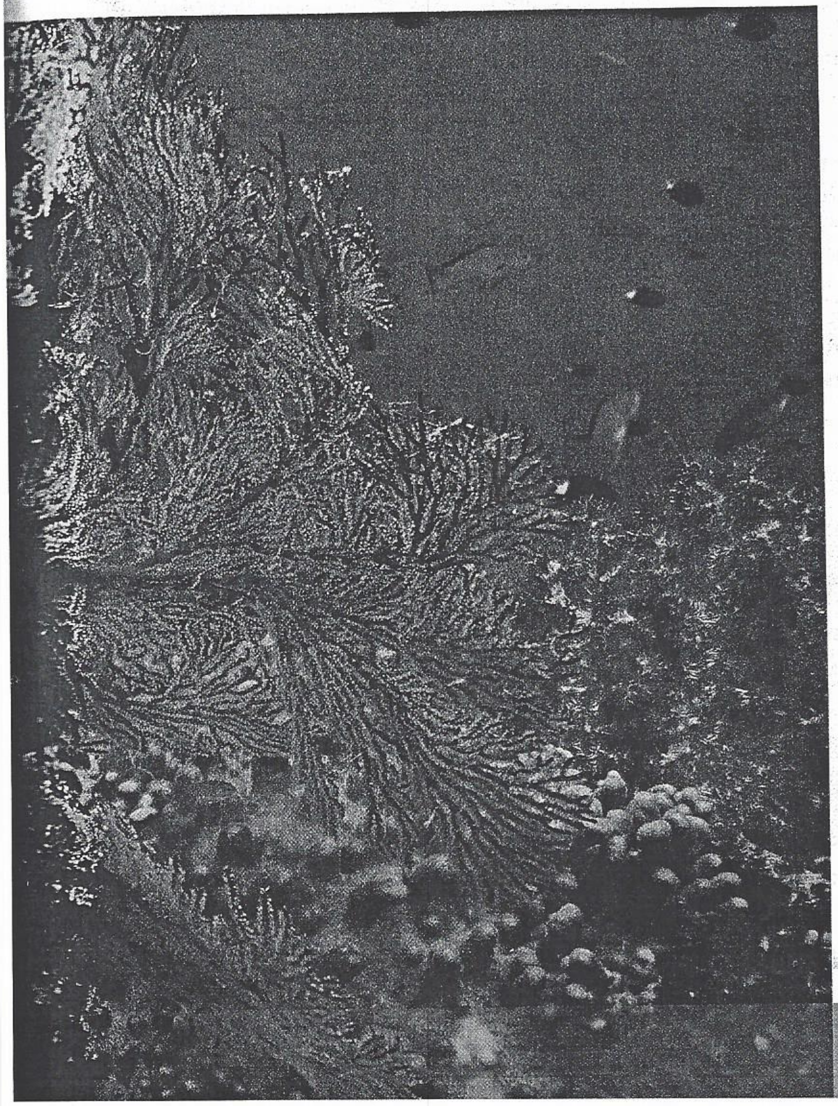
? ***Will rising atmospheric CO_2 harm coral reefs?***

Why do we begin our study of biology with a chapter on chemistry? Well, chemistry is the basis of life—it explains how elements combine into the compounds that make up your body and the bodies of all other living organisms and how chemical reactions underlie the functions of all cells.

Life and its chemistry are tied to water. Life began in water and evolved there for 3 billion years before spreading onto land. And all life, even land-dwelling life, is still dependent on water. Your

cells are about 75% water, and that is where the chemical reactions of your body take place. What properties of the simple water molecule make it so indispensable to life on Earth? You'll find out in this chapter.

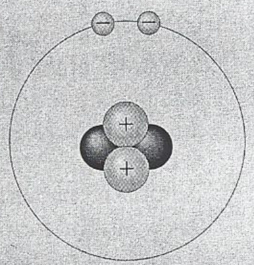
This chapter will also make connections to one of the main themes in biology—the organization of life into a hierarchy of structural levels, with new properties emerging at each successive level (see Chapter 1). You will see that emergent properties are apparent even at the lowest levels of biological organization—the ordering of atoms into molecules and the interactions of those molecules. Thus we begin our study of biology with some basic concepts of chemistry that will apply throughout our study of life.



BIG IDEAS

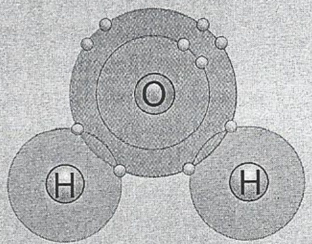
Elements, Atoms, and Compounds
(2.1–2.4)

Living organisms are made of atoms of certain elements, mostly combined into compounds.



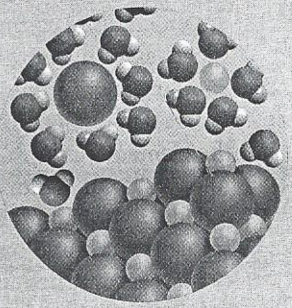
Chemical Bonds
(2.5–2.9)

The structure of an atom determines what types of bonds it can form with other atoms.



Water's Life-Supporting Properties
(2.10–2.16)

The unique properties of water derive from the polarity and hydrogen bonding of water molecules.



Elements, Atoms, and Compounds

2.1 Organisms are composed of elements, in combinations called compounds

You and all things around you are made of matter—the physical "stuff" of the universe. **Matter** is defined as anything that occupies space and has mass. (In everyday language, we can think of mass as an object's weight.) Matter is found on Earth in three physical states: solid, liquid, and gas. Water is a rare example of matter that exists in the natural environment in all three physical forms: as ice, liquid water, and water vapor.

Types of matter as diverse as water, rocks, air, and biology students are all composed of chemical elements. An **element** is a substance that cannot be broken down to other substances by ordinary chemical means. Today, chemists recognize 92 elements that occur in nature; gold, copper, carbon, and oxygen are some examples. Chemists have also made a few dozen synthetic elements. Each element has a symbol, the first letter or two of its English, Latin, or German name. For example, the symbol for sodium, Na, is from the Latin word *natrium*; the symbol O comes from the English word *oxygen*.

A **compound** is a substance consisting of two or more different elements combined in a fixed ratio. Compounds are much more common than pure elements. In fact, few elements exist in a pure state in nature.

Many compounds consist of only two elements; for instance, table salt (sodium chloride, NaCl) has equal parts of the elements sodium (Na) and chlorine (Cl). Pure sodium is a metal and pure chlorine is a poisonous gas. Chemically combined, however, they form an edible compound (**Figure 2.1**). Hydrogen (H) and oxygen (O) are elements that typically exist as gases. Chemically combined in a ratio of 2:1, however, they form the most abundant compound on the surface of Earth—water (H_2O). These are simple examples of organized matter having emergent properties: A compound has characteristics different from those of its elements.

Most of the compounds in living organisms contain at least three or four elements. Sugar, for example, is formed of carbon (C), hydrogen, and oxygen. Proteins are compounds containing carbon, hydrogen, oxygen, nitrogen (N), and a small amount of sulfur (S). Different arrangements of the atoms of these elements give rise to the unique properties of each compound.

How many of the 92 natural elements are essential for life? The requirements are similar among organisms, but there is some variation. For example, humans need 25 elements, but plants need only 17. Four elements—oxygen, carbon, hydrogen, and nitrogen—make up about 96% of all living matter. These elements are the main ingredients of biological molecules. As you can see in **Table 2.1**, which lists the 25 elements found in humans, calcium (Ca), phosphorus (P), potassium (K), sulfur, sodium, chlorine, and magnesium (Mg) account for most of the remaining 4% of your body. These elements are involved in such important functions as bone formation (calcium and phosphorus) and nerve signaling (potassium, sodium, calcium, and chlorine).

The **trace elements** listed at the bottom of the table are essential for humans, but only in minute quantities. Some trace elements, such as iron (Fe), are needed by all forms of life. Iron makes up only about 0.004% of your body weight but is vital for energy processing and for transporting oxygen in your blood. Other trace elements are required only by certain species. For example, iodine is an essential element only for vertebrates—animal with backbones, which, of course, includes you. We explore the importance of trace elements to your health next.

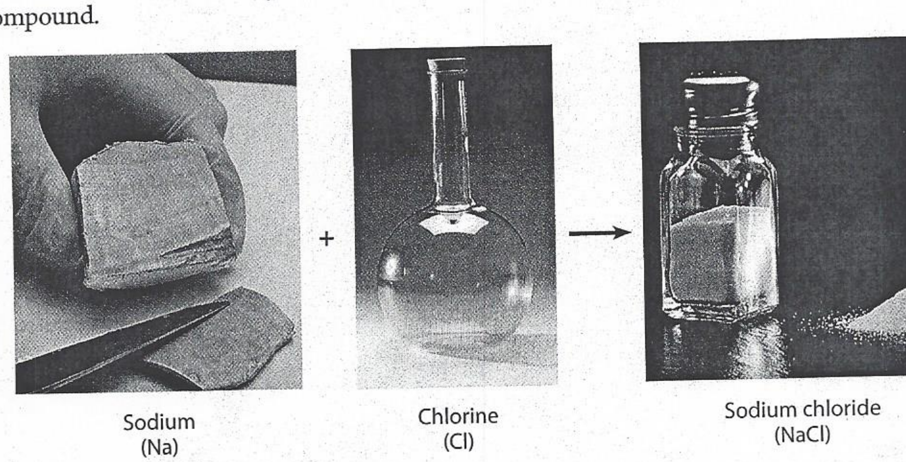
TABLE 2.1 | ELEMENTS IN THE HUMAN BODY

Element	Symbol	Percentage of Body Weight (Including Water)	
Oxygen	O	65.0	96.3%
Carbon	C	18.5	
Hydrogen	H	9.5	
Nitrogen	N	3.3	
Calcium	Ca	1.5	3.7%
Phosphorus	P	1.0	
Potassium	K	0.4	
Sulfur	S	0.3	
Sodium	Na	0.2	
Chlorine	Cl	0.2	
Magnesium	Mg	0.1	

Trace elements, less than 0.01% of human body weight: Boron (B), chromium (Cr), cobalt (Co), copper (Cu), fluorine (F), iodine (I), iron (Fe), manganese (Mn), molybdenum (Mo), selenium (Se), silicon (Si), tin (Sn), vanadium (V), zinc (Zn)

? Explain how table salt illustrates the theme of emergent properties.

The elements that make up the edible crystals of table salt, sodium and chlorine, are in pure form a metal and a poisonous gas.



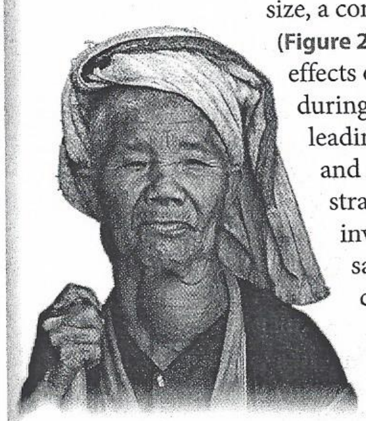
▲ **Figure 2.1** The emergent properties of the edible compound sodium chloride

2.2 Trace elements are common additives to food and water

CONNECTION

Trace elements are required in very small quantities, but, in some cases, even those small requirements are difficult to fulfill.

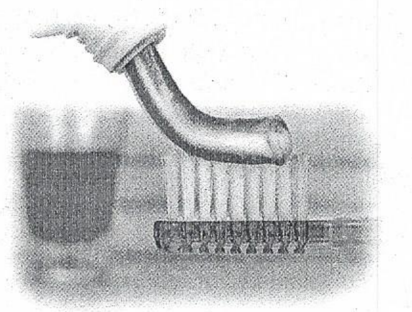
Iodine is an essential ingredient of a hormone produced by the thyroid gland, which is located in your neck. You need to ingest only a tiny speck of iodine each day, about 0.15 milligram (mg). An iodine deficiency in the diet causes the thyroid gland to grow to abnormal size, a condition called goiter (**Figure 2.2A**). The most serious effects of iodine deficiency take place during fetal development and childhood, leading to miscarriages, poor growth, and mental impairment. A global strategy to eliminate iodine deficiency involves universal iodization of all salt used for human and animal consumption. Unfortunately, about 30% of global households still do not have access to iodized salt, and an estimated 2 billion people are still at risk of iodine deficiency. Seafood, kelp, dairy products, and dark, leafy greens are good natural sources. Thus, deficiencies are often found in inland regions, especially in areas where the soil is lacking in iodine. Although most common in developing nations, iodine deficiencies may also result from excessive consumption of highly processed foods (which often use non-iodized salt) and low-salt diets intended to lower the risk of cardiovascular disease.



▲ **Figure 2.2A** Goiter, a symptom of iodine deficiency, in a Burmese woman

Iodine is just one example of a trace element added to food or water to improve health. For more than 60 years, the American Dental Association has supported fluoridation of community drinking water as a public health measure. Fluoride is a form of fluorine (F), an element in Earth's crust that is found in small amounts in all water sources. In many communities, fluoride is added during the municipal water treatment process to raise levels to a concentration that can reduce tooth decay. If you mostly drink bottled water, however, your fluoride intake may be reduced, although some bottled water now contains added fluoride. Fluoride is also frequently added to dental products, such as toothpaste and mouthwash (**Figure 2.2B**).

Chemicals are added to food to help preserve it, make it more nutritious, or simply make it look better. Read the nutrition facts label from the side of the cereal box in **Figure 2.2C** to see a familiar example of how foods are fortified with mineral elements. Iron, for example, is commonly added to foods. (You can actually see that iron has been added to a fortified cereal by crushing the cereal and then stirring a magnet through it.) Also note that the nutrition facts label lists numerous vitamins that are added to improve the nutritional value of the cereal. For instance, the cereal in this example supplies 10% of the recommended daily value for vitamin A. Vitamins consist of more than one element and are examples of compounds.

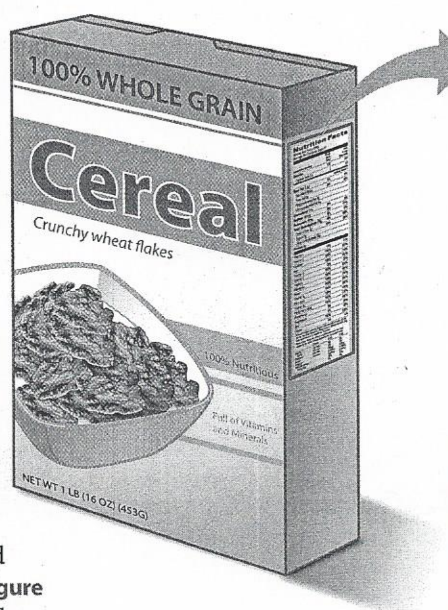


◀ **Figure 2.2B** Mouthwash and toothpaste with added fluoride

In the next module, we explore the chemical properties of elements and how the structure of an atom—the smallest unit of an element—determines those properties.

? **In addition to iron, what other trace elements are found in the cereal in Figure 2.2C? Does one serving provide the total daily amount needed of these elements?**

● Zinc and copper: one serving provides 100% of the zinc but only 4% of the copper needed in a day.



Nutrition Facts
Serving Size ¾ cup (30g)
Servings Per Container about 17

Amount Per Serving	Whole Grain Cereal	with ½ cup skim milk
Calories	100	140
Calories from Fat	5	10
		% Daily Value**
Total Fat 0.5g*	**1%**	**1%**
Saturated Fat 0g	**0%**	**0%**
Trans Fat 0g		
Polyunsaturated Fat 0g		
Monounsaturated Fat 0g		
Cholesterol 0mg	**0%**	**1%**
Sodium 135mg	**6%**	**9%**
Potassium 125mg	**4%**	**10%**
Total Carbohydrate 23g	**8%**	**10%**
Dietary Fiber 3g	**10%**	**10%**
Sugars 5g		
Other Carbohydrate 15g		
Protein 2g		
Vitamin A	10%	15%
Vitamin C	100%	100%
Calcium	100%	110%
Iron	100%	100%
Vitamin D	10%	25%
Vitamin E	100%	100%
Thiamin	100%	100%
Riboflavin	100%	110%
Niacin	100%	100%
Vitamin B_6	100%	100%
Folic Acid	100%	100%
Vitamin B_{12}	100%	110%
Pantothenic Acid	100%	100%
Phosphorus	8%	20%
Magnesium	6%	10%
Zinc	100%	100%
Copper	4%	4%

*Amount in cereal. A serving of cereal plus skim milk provides 1g total fat, less than 5mg cholesterol, 260mg sodium, 290mg potassium, 29g total carbohydrate (11g sugars) and 7g protein.

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

	Calories	2,000	2,500
Total Fat	Less than	65g	80g
Sat Fat	Less than	20g	25g
Cholesterol	Less than	300mg	300mg
Sodium	Less than	2,400mg	2,400mg
Potassium		3,500mg	3,500mg
Total Carbohydrate		300g	375g
Dietary Fiber		25g	30g

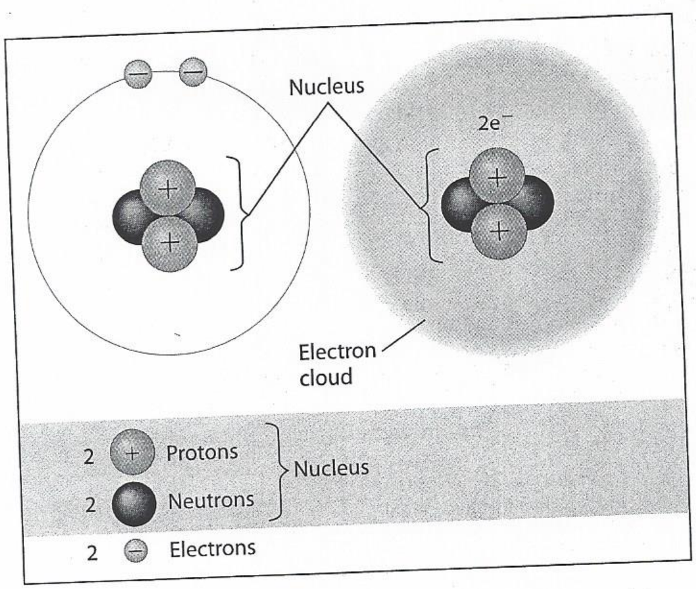
▲ **Figure 2.2C** Nutrition facts from a fortified cereal

2.3 Atoms consist of protons, neutrons, and electrons

Each element has its own type of atom, which is different from the atoms of other elements. An **atom**, named from a Greek word meaning "indivisible," is the smallest unit of matter that still retains the properties of an element. Atoms are so small that it would take about a million of them to stretch across the period printed at the end of this sentence.

Subatomic Particles Physicists have split the atom into more than a hundred types of subatomic particles. However, only three kinds of particles are relevant here. A **proton** is a subatomic particle with a single positive electrical charge (+). An **electron** is a subatomic particle with a single negative charge (−). A **neutron**, as its name implies, is electrically neutral (has no charge).

Figure 2.3 shows two very simple models of an atom of the element helium (He), the "lighter-than-air" gas that makes balloons rise. Notice that two protons (⊕) and two neutrons (●) are tightly packed in the atom's central core, or **nucleus**. Two electrons (⊖) form a sort of cloud of negative charge around the nucleus. The attraction between the negatively charged electrons and the positively charged protons holds the electrons near the nucleus. The left-hand model shows the two electrons on a circle around the nucleus. The right-hand model, slightly more realistic, shows a spherical cloud of negative charge created by the two rapidly moving electrons. Neither model is drawn to scale. In real atoms, the electrons are very much smaller than the protons and neutrons, and the electron cloud is very much bigger compared to the nucleus. Imagine that this atom was the size of a baseball stadium: The nucleus would be the size of a pea in center field, and the electrons would be like two tiny gnats buzzing around the stadium.



▲ **Figure 2.3** Two models of a helium atom. (Note that these models are not to scale; they greatly overestimate the size of the nucleus in relation to the electron cloud.)

Atomic Number and Mass Number So what makes the atoms of different elements different? All atoms of a particular element have the same unique number of protons. This number is the element's **atomic number**. Thus, an atom of helium, with 2 protons, has an atomic number of 2. Unless otherwise indicated, an atom has an equal number of protons and electrons, and thus its net electrical charge is 0 (zero).

What other numbers are associated with an atom? An atom's **mass number** is the sum of the number of protons and neutrons in its nucleus. For helium, the mass number is 4. The mass of a proton and the mass of a neutron are almost identical and are expressed in a unit of measurement called the dalton. Protons and neutrons each have masses close to 1 dalton. An electron has only about 1/2,000 the mass of a proton, so it contributes very little to an atom's mass. Thus, an atom's **atomic mass** (or weight) is approximately equal to its mass number—the sum of its protons and neutrons—in daltons.

Isotopes All atoms of an element have the same atomic number, but some atoms of that element may differ in mass number. The different **isotopes** of an element have the same number of protons and behave identically in chemical reactions, but they have different numbers of neutrons. **Table 2.3** shows the numbers of subatomic particles in the three isotopes of carbon. Note that carbon's atomic number is 6—all of its atoms have 6 protons. Carbon-12 (named for its mass number), with 6 neutrons, accounts for about 99% of naturally occurring carbon. Most of the remaining 1% consists of carbon-13, with a mass number of 13 and thus 7 neutrons. A third isotope, carbon-14, with 8 neutrons, occurs in minute quantities. Of course, all three isotopes have 6 protons—otherwise, they would not be carbon.

Both carbon-12 and carbon-13 are stable isotopes, meaning that their nuclei remain intact more or less forever. The isotope carbon-14, on the other hand, is unstable, or radioactive. A **radioactive isotope** is one in which the nucleus decays spontaneously, giving off particles and energy. Radiation from decaying isotopes can damage cellular molecules and thus can pose serious risks to living organisms. But radioactive isotopes can be helpful, as in their use in dating fossils (see Module 15.5). They are also used in biological research and medicine, as we see next.

TABLE 2.3 | ISOTOPES OF CARBON

	Carbon-12		Carbon-13		Carbon-14	
Protons	6	Mass number 12	6	Mass number 13	6	Mass number 14
Neutrons	6		7		8	
Electrons	6		6		6	

? **A nitrogen atom has 7 protons, and its most common isotope has 7 neutrons. A radioactive isotope of nitrogen has 9 neutrons. What is the atomic number and mass number of this radioactive nitrogen?**

Atomic number = 7; mass number = 16

2.4 Radioactive isotopes can help or harm us

CONNECTION

Living cells cannot readily distinguish between isotopes of the same element. Consequently, organisms take up and use compounds containing radioactive isotopes in the usual way. Because radioactivity is easily detected and measured by instruments, radioactive isotopes are useful as tracers—biological spies, in effect—for monitoring the fate of atoms in living organisms.

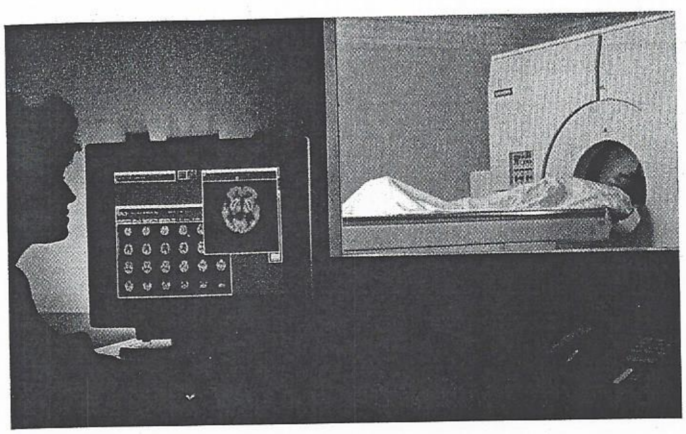
Basic Research Biologists often use radioactive tracers to follow molecules as they undergo chemical changes in an organism. For example, researchers have used carbon dioxide (CO_2) containing the radioactive isotope carbon-14 to study photosynthesis. Using sunlight to power the conversion, plants take in CO_2 from the air and use it to make sugar molecules. Radioactively labeled CO_2 has enabled researchers to trace the sequence of molecules made by plants in the chemical route from CO_2 to sugar.

Medical Diagnosis and Treatment Radioactive isotopes may also be used to tag chemicals that accumulate in specific areas of the body, such as phosphorus in bones. After injection of such a tracer, a special camera produces an image of where the radiation collects. In most diagnostic uses, the patient receives only a tiny amount of an isotope.

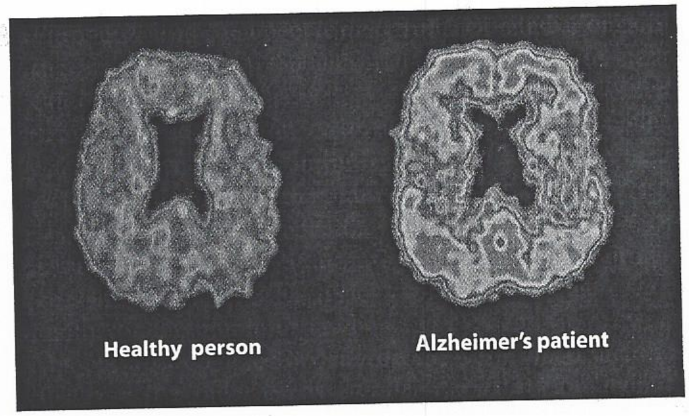
Sometimes radioactive isotopes are used for treatment. As you learned in Module 2.2, the body uses iodine to make a thyroid hormone. Because radioactive iodine accumulates in the thyroid, it can be used to kill cancer cells there.

Substances that the body metabolizes, such as glucose or oxygen, may also be labeled with a radioactive isotope. **Figure 2.4A** shows a patient being examined by a PET (positron-emission tomography) scanner, which can produce three-dimensional images of areas of the body with high metabolic activity. PET is useful for diagnosing certain heart disorders and cancers and for basic research on the brain.

The early detection of Alzheimer's disease may be a new use for such techniques. This devastating illness gradually destroys a person's memory and ability to think. As the disease progresses, the brain becomes riddled with deposits (plaques) of a protein called beta-amyloid. Researchers have identified



▲ **Figure 2.4A** Technician monitoring the output of a PET scanner



▲ **Figure 2.4B** PET images of brains of a healthy person (left) and a person with Alzheimer's disease (right). Red and yellow colors indicate high levels of PIB bound to beta-amyloid plaques.

a protein molecule called PIB that binds to beta-amyloid. PIB contains a radioactive isotope that can be detected on a PET scan. **Figure 2.4B** shows PET images of the brains of a healthy person (left) and a person with Alzheimer's (right) injected with PIB. Notice that the brain of the Alzheimer's patient has high levels of PIB (red and yellow areas), whereas the unaffected person's brain has lower levels (blue). New therapies are focused on limiting the production of beta-amyloid or clearing it from the brain. A diagnostic test using PIB would allow researchers to monitor the effectiveness of new drugs in people living with the disease.

Dangers Although radioactive isotopes have many beneficial uses, uncontrolled exposure to them can harm living organisms by damaging molecules, especially DNA. The particles and energy thrown off by radioactive atoms can break chemical bonds and also cause abnormal bonds to form. The explosion of a nuclear reactor in Chernobyl, Ukraine, in 1986 released large amounts of radioactive isotopes into the environment, which drifted over large areas of Russia, Belarus, and Europe. A few dozen people died from acute radiation poisoning, and more than 100,000 people were evacuated from the immediate area. Increased rates of thyroid cancer in children exposed to the radiation have been reported. Likewise, scientists will carefully monitor the long-term health consequences of the 2011 post-tsunami Fukushima nuclear disaster in Japan.

Natural sources of radiation can also pose a threat. Radon, a radioactive gas, may be a cause of lung cancer. Radon can contaminate buildings in regions where underlying rocks naturally contain uranium, a radioactive element. Homeowners can buy a radon detector or hire a company to test their home to ensure that radon levels are safe. If levels are found to be unsafe, technology exists to remove radon from homes.

? **Why are radioactive isotopes useful as tracers in research on the chemistry of life?**

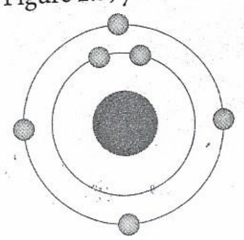
● Organisms incorporate radioactive isotopes of an element into their molecules, and researchers can use special scanning devices to detect the presence of these isotopes in biological pathways or locations in the body.

Chemical Bonds

2.5 The distribution of electrons determines an atom's chemical properties

To understand how atoms interact with each other, we need to explore atomic structure further. Of the three subatomic particles—protons, neutrons, and electrons—only electrons are directly involved in the chemical activity of an atom.

If you glance back to the model of the helium atom in Figure 2.3, you see that its 2 electrons are shown together on a circle around the nucleus. But where should the electrons be shown in an atom with more than 2 electrons—say, in carbon, whose atomic number is 6? As you see in **Figure 2.5A**, 2 electrons are still shown on an inner circle, but the next 4 are placed on a larger outside circle. It turns out that electrons can be located in different **electron shells**, each with a characteristic distance from the nucleus. Depending on an element's atomic number, an atom may have one, two, or more electron shells.



▲ **Figure 2.5A** An electron distribution model of carbon

Figure 2.5B is an abbreviated version of the periodic table of the elements (see Appendix 2 for the complete table). The figure shows the distribution of electrons for the first 18 elements, arranged in rows according to the number of electron shells (one, two, or three). Within each shell, electrons travel in different *orbitals*, which are discrete volumes of space in which electrons are most likely to be found. Each orbital can hold a maximum of 2 electrons. The first electron shell has only one orbital and can hold only 2 electrons. Thus, hydrogen and helium are the only elements in the first row. For the second and third rows, the outer shell has four orbitals and can hold up to 8 electrons (four pairs).

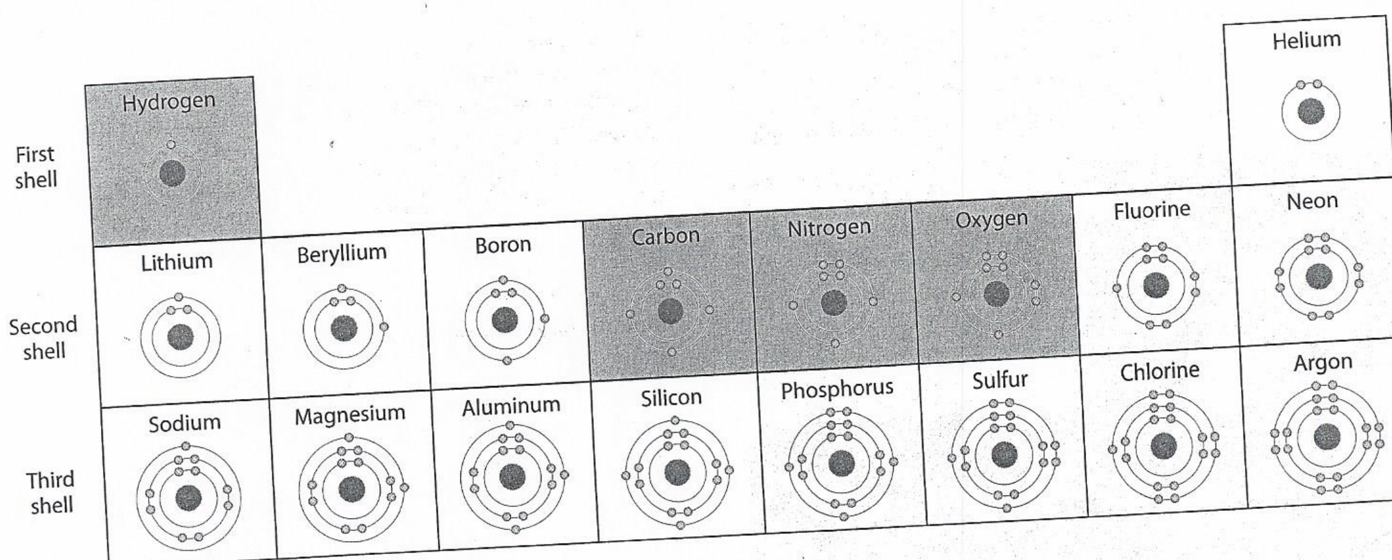
It is the number of electrons present in the outermost shell, called the valence shell, that determines the chemical properties of an atom. Atoms whose outer shells are not full tend to interact with other atoms in ways that enable them to complete or fill their valence shells.

Look at the electron shells of the atoms of the four elements that are the main components of biological molecules (highlighted in green in Figure 2.5B). Because their outer shells are incomplete, all these atoms react readily with other atoms. The hydrogen atom has only 1 electron in its single electron shell, which can accommodate 2 electrons. Atoms of carbon, nitrogen, and oxygen also have unpaired electrons and incomplete shells. In contrast, the helium atom has a first-level shell that is full with 2 electrons. Neon and argon also have full outer shells. As a result, these elements are chemically inert (unreactive).

How do chemical interactions between atoms enable them to fill their outer electron shells? When two atoms with incomplete outer shells react, each atom will share, donate, or receive electrons, so that both partners end up with completed outer shells. These interactions usually result in atoms staying close together, held by attractions known as **chemical bonds**. In the next two modules, we look at two important types of chemical bonds.

? **How many electrons and electron shells does a sodium atom have? How many electrons are in its valence shell?**

● 11 electrons; 3 electron shells; 1 electron in the outer shell



▲ **Figure 2.5B** The electron distribution diagrams of the first 18 elements in the periodic table

Try This As you read from left to right across each row, describe how the number of electrons changes. Note that the electrons don't pair up until all orbitals have at least one electron.

2.6 Covalent bonds join atoms into molecules through electron sharing

In a **covalent bond**, two atoms, each with an unpaired electron in its outer shell, actually *share* a pair of electrons. Sharing one or more pairs of electrons enables atoms to complete their outer shells. Atoms held together by covalent bonds form a **molecule**. For example, a covalent bond connects two hydrogen atoms in a molecule of the gas H_2, and a covalent bond connects each of two hydrogen atoms to an oxygen atom in a molecule of water (H_2O).

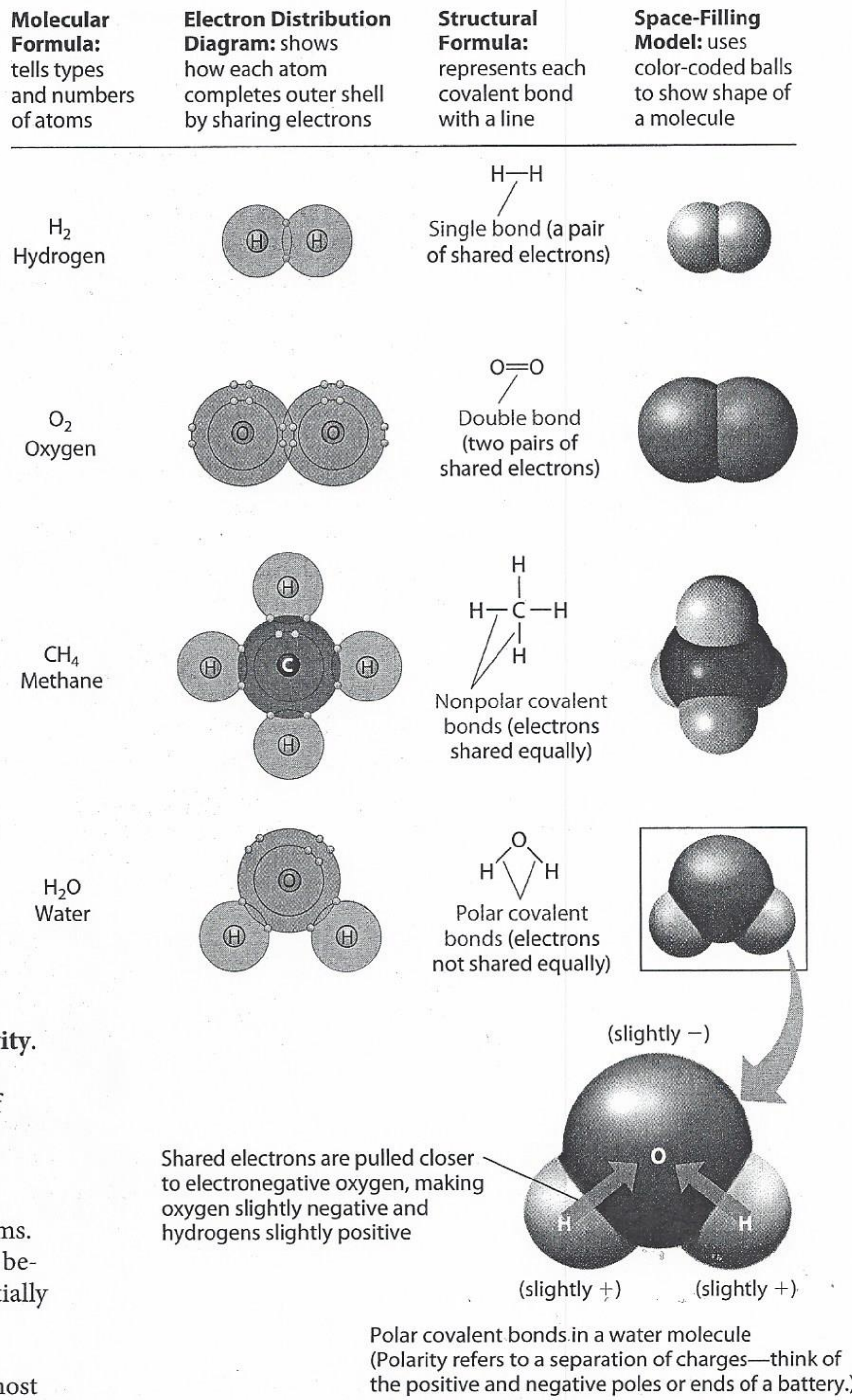
How many covalent bonds can an atom form? It depends on the number of additional electrons needed to fill its valence shell. This number is called the valence, or bonding capacity, of an atom. Look back at the electron distribution diagrams in Figure 2.5B and see if you can determine how many covalent bonds hydrogen, oxygen, nitrogen, and carbon can form.

Figure 2.6 shows how molecules can be represented in several different ways. Let's see what we can learn from this figure. As you will notice in the electron distribution diagram, the hydrogen atoms in H_2 are held together by a pair of shared electrons. But two atoms can share more than just one pair of electrons. In an oxygen molecule (O_2), for example, the two oxygen atoms share two pairs of electrons, forming a double bond. A double bond is indicated in a structural formula by a pair of lines.

H_2 and O_2 are molecules composed of only one element. Methane (CH_4) and water (H_2O) are compounds. Methane is a major component of natural gas. As shown in Figure 2.6, it takes four hydrogen atoms to satisfy carbon's valence of 4.

Atoms in a molecule are in a constant tug-of-war for the shared electrons of their covalent bonds. An atom's attraction for shared electrons is called its **electronegativity**. The more electronegative an atom, the more strongly it pulls shared electrons toward its nucleus. In molecules of only one element, such as H_2 and O_2, the two identical atoms exert an equal pull on the electrons. The bonds in such molecules are said to be **nonpolar covalent bonds** because the electrons are shared equally between the atoms. Compounds such as methane also have nonpolar bonds, because the atoms of carbon and hydrogen are not substantially different in electronegativity.

Water, on the other hand, is composed of atoms with quite different electronegativities. Oxygen is one of the most electronegative of the elements. As indicated by the arrows in the blowup of a water molecule in Figure 2.6, oxygen attracts the shared electrons in H_2O much more strongly than does hydrogen, so that the electrons spend more time near the oxygen atom than near the hydrogen atoms. This unequal sharing of electrons produces a **polar covalent bond**. In a polar covalent bond, the pulling of shared, negatively charged electrons closer to the more electronegative atom makes that atom partially negative and the other atom partially positive. Thus, in H_2O, the oxygen atom actually has a slight negative charge and each hydrogen atom a slight positive charge.



▲ **Figure 2.6** Alternative ways to represent four common molecules

In some cases, two atoms are so unequal in their attraction for electrons that the more electronegative atom strips an electron completely away from its partner, as we see next.

? **What is chemically nonsensical about this structure?**

H—C=C—H

● Each C has only three bonds instead of the four required by its valence.

2.7 Ionic bonds are attractions between ions of opposite charge

Table salt is an example of how the transfer of electrons can bond atoms together. **Figure 2.7A** shows how a sodium atom (Na) and a chlorine atom (Cl) can form the compound sodium chloride (NaCl). Notice that sodium has only 1 electron in its outer shell, whereas chlorine has 7. When these atoms interact, the sodium atom transfers its single outer electron to chlorine. Sodium now has only two shells, the second shell having a full set of 8 electrons. When chlorine strips away sodium's electron, its own outer shell is now full with 8 electrons.

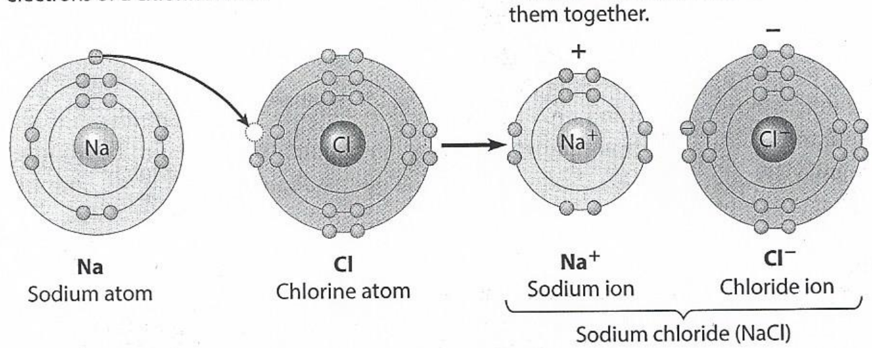
But how does this electron transfer result in an ionic bond between sodium and chlorine? Remember that electrons are negatively charged particles. The transfer of an electron moves one unit of negative charge from one atom to the other. Sodium, with 11 protons but now only 10 electrons, has a net electrical charge of 1+. Chlorine, having gained an extra electron, now has 18 electrons but only 17 protons, giving it a net electrical charge of 1−. In each case, an atom has become an **ion**—an atom or molecule with an electrical charge resulting from a gain or loss of one or more electrons. (Note that the names of negatively charged ions often end in *-ide*, such as *chloride*.) Two ions with opposite charges attract each other. When the attraction holds them together, it is called an **ionic bond.** The resulting compound, in this case NaCl, is electrically neutral.

Sodium chloride is a familiar type of **salt**, a synonym for an ionic compound. Salts often exist as crystals in nature. **Figure 2.7B** shows the ions Na^+ and Cl^- in a crystal of sodium chloride. An NaCl crystal can be of any size (there is no fixed number of ions), but sodium and chloride ions are always present in a 1:1 ratio. The ratio of ions can differ in the various kinds of salts.

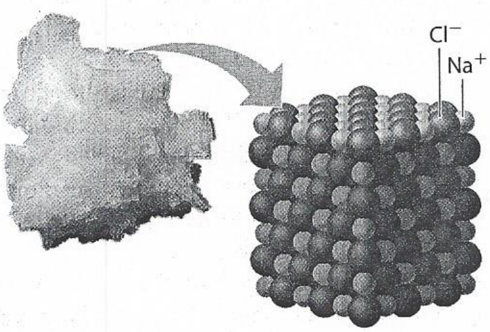
The environment affects the strength of ionic bonds. In a dry salt crystal, the bonds are so strong that it takes a hammer and chisel to break enough of them to crack the crystal. If the same salt crystal is placed in water, however, the ionic bonds break when the ions interact with water molecules and the salt dissolves, as we'll discuss in Module 2.13. Most drugs are manufactured as salts because they are quite stable when dry but can dissolve easily in water.

? Explain what holds together the ions in a crystal of table salt (NaCl).

Opposite charges attract. The positively charged sodium ions (Na^+) and the negatively charged chloride ions (Cl^-) are held together by ionic bonds, attractions between oppositely charged ions.



▲ **Figure 2.7A** Formation of an ionic bond, producing sodium chloride



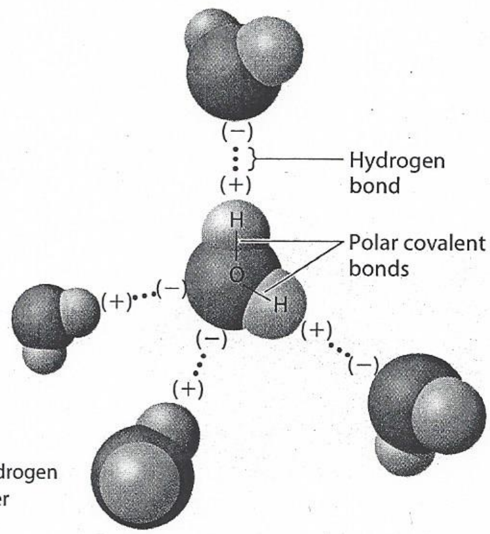
▲ **Figure 2.7B** A crystal of sodium chloride

2.8 Hydrogen bonds are weak bonds important in the chemistry of life

In living organisms, most of the strong chemical bonds are covalent, linking atoms to form a cell's molecules. But crucial to the functioning of a cell are weaker bonds within and between molecules, such as the ionic bonds we just discussed. One of the most important types of weak bonds is the **hydrogen bond**, which is best illustrated with water molecules.

As you saw in Figure 2.6, the hydrogen atoms of a water molecule are attached to oxygen by polar covalent bonds. Because of these polar bonds and the wide V shape of the molecule, water is a **polar molecule**—that is, it has an unequal distribution of charges. It is slightly negative at the oxygen end of the molecule (the point of the V) and slightly positive at each of the two hydrogen ends. This partial positive charge allows each hydrogen to be attracted to—in a sense, to "flirt" with—a nearby atom (often an oxygen or nitrogen) that has a partial negative charge.

Figure 2.8, on the next page, illustrates how these weak bonds form between water molecules. They are called



▶ **Figure 2.8** Hydrogen bonds between water molecules

Try This Describe polar covalent bonds and hydrogen bonds and explain how they are related.

hydrogen bonds because the positively charged atom in this type of attraction is always a hydrogen atom. As Figure 2.8 shows, each hydrogen atom of a water molecule can form a hydrogen bond (depicted by dotted lines) with a nearby partially negative oxygen atom of another water molecule. And the negative (oxygen) pole of a water molecule can form hydrogen bonds to two hydrogen atoms. Thus, each water molecule can hydrogen-bond to as many as four partners.

You will learn later how hydrogen bonds help to create a protein's shape (and thus its function) and hold the two strands of a DNA molecule together (see Chapter 3). Later in this chapter, we explore how water's polarity and hydrogen bonds give it unique, life-supporting properties. But first we discuss how the making and breaking of bonds change the composition of matter.

? **What enables neighboring water molecules to hydrogen-bond to one another?**

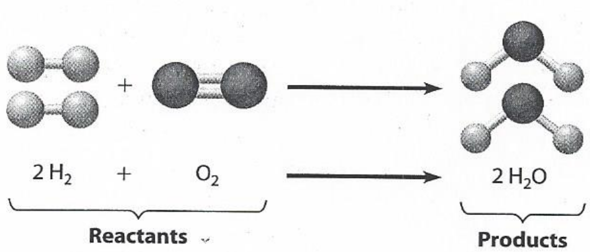
● The molecules are polar, with each positive end (hydrogen end) of one molecule attracted to the negative end (oxygen end) of another molecule.

2.9 Chemical reactions make and break chemical bonds

Your cells are constantly rearranging molecules in **chemical reactions**—breaking existing chemical bonds and forming new ones. A simple example of a chemical reaction is the reaction between hydrogen gas and oxygen gas that forms water (this is an explosive reaction, which, fortunately, does not occur in your cells):

$$2\ H_2 + O_2 \rightarrow 2\ H_2O$$

In this case, two molecules of hydrogen (2 H_2) react with one molecule of oxygen (O_2) to produce two molecules of water (2 H_2O). The arrow in the equation indicates the conversion of the starting materials, called the **reactants**, to the **product**, the material resulting from the chemical reaction. Notice that the same *numbers* of hydrogen and oxygen atoms appear on the left and right sides of the arrow, although they are grouped differently. Chemical reactions do not create or destroy matter; they only rearrange it in various ways. As shown in **Figure 2.9**, the covalent bonds (represented here as white "sticks" between atoms) holding hydrogen atoms together in H_2 and holding oxygen atoms together in O_2 are broken, and new bonds are formed to yield the H_2O product molecules.



▲ **Figure 2.9** Breaking and making of bonds in a chemical reaction

Organisms cannot make water from H_2 and O_2, but they do carry out a great number of chemical reactions that rearrange matter in significant ways. Let's examine a chemical reaction that is essential to life on Earth: photosynthesis. The raw materials of photosynthesis are carbon dioxide (CO_2), which is taken from the air, and water (H_2O), which plants absorb from the soil. Within green plant cells, sunlight powers the conversion of these reactants to the sugar product glucose ($C_6H_{12}O_6$) and oxygen (O_2), a by-product that the plant releases into the air. The following chemical shorthand summarizes the process:

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

Although photosynthesis is actually a sequence of many chemical reactions, we see that we end up with the same number and kinds of atoms we started with. Matter has simply been rearranged, with an input of energy provided by sunlight.

Your body routinely carries out thousands of chemical reactions. These reactions take place in the watery environment of your cells. We look at the life-supporting properties of water next.

? **Fill in the blanks with the correct numbers in the following chemical process:**

$$C_6H_{12}O_6 + \_\_O_2 \rightarrow \_\_CO_2 + \_\_H_2O$$

What process do you think this reaction represents? (*Hint:* Think about how your cells use these reactants to produce energy.)

● $C_6H_{12}O_6 + 6\ O_2 \rightarrow 6\ CO_2 + 6\ H_2O$; the breakdown of sugar in the presence of oxygen to carbon dioxide and water, with the release of energy that the cell can use

Water's Life-Supporting Properties

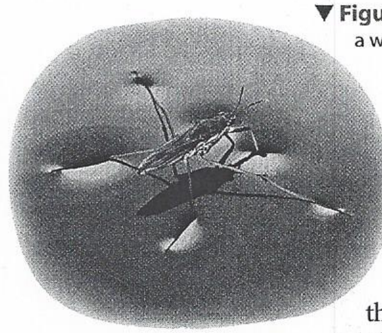
2.10 Hydrogen bonds make liquid water cohesive

We can trace water's life-supporting properties to the structure and interactions of its molecules—their polarity and resulting hydrogen bonding between molecules (review Figure 2.8).

Hydrogen bonds between molecules of liquid water last for only a few trillionths of a second, yet at any instant, many molecules are hydrogen-bonded to others. This tendency of molecules of the same kind to stick together, called **cohesion**, is much stronger for water than for most other liquids. The cohesion of water is important in the living world. Trees, for example, depend on cohesion to help transport water and nutrients from their roots to their leaves. The evaporation of water from a leaf exerts a pulling force on water within the veins of the leaf. Because of cohesion, the force is relayed all the way down to the roots. **Adhesion**, the clinging of one substance to another, also plays a role. The adhesion of water to the cell walls of a plant's thin veins helps counter the downward pull of gravity.

Related to cohesion is **surface tension**, a measure of how difficult it is to stretch or break the surface of a liquid. Hydrogen bonds give water unusually high surface tension, making it behave as though it were coated with an invisible film. You can observe the surface tension of water by slightly overfilling a glass; the water will stand above the rim. The water strider in **Figure 2.10** takes advantage of the high surface tension of water to "stride" across ponds without breaking the surface.

▼ **Figure 2.10** Surface tension allowing a water strider to walk on water



? **After a hard workout, you may notice "beads" of sweat on your face. Can you explain what holds the sweat in droplet form?**

● The cohesion of water molecules and its high surface tension hold water in droplets. The adhesion of water to your skin helps hold the beads in place.

2.11 Water's hydrogen bonds moderate temperature

Thermal energy is the energy associated with the random movement of atoms and molecules. Thermal energy in transfer from a warmer to a cooler body of matter is defined as **heat**. **Temperature** measures the intensity of heat—that is, the *average* speed of molecules in a body of matter. If you have ever burned your finger on a metal pot while waiting for the water in it to boil, you know that water heats up much more slowly than metal. In fact, because of hydrogen bonding, water has a stronger resistance to temperature change than most other substances.

▲ **Figure 2.11** Sweating as a mechanism of evaporative cooling

Heat must be absorbed to break hydrogen bonds, and heat is released when hydrogen bonds form. To raise the temperature of water, hydrogen bonds between water molecules must be broken before the molecules can move faster. Thus, water absorbs a large amount of heat (much of it used to disrupt hydrogen bonds) while warming up only a few degrees. Conversely, when water cools, water molecules slow down and more hydrogen bonds form, releasing a considerable amount of heat.

Earth's giant water supply moderates temperatures, helping to keep temperatures within limits that permit life. Oceans, lakes, and rivers store a huge amount of heat from the sun during warm periods. Heat given off from gradually cooling water warms the air. That's why coastal areas generally have milder climates than inland regions. Water's resistance to temperature change also stabilizes ocean temperatures, creating a favorable environment for marine life. Because water accounts for approximately 66% of your body weight, it also helps moderate your temperature.

When a substance evaporates (changes physical state from a liquid to a gas), the surface of the liquid that remains behind cools down. This **evaporative cooling** occurs because the molecules with the greatest energy (the "hottest" ones) leave. It's as if the 10 fastest runners on the track team left school, lowering the average speed of the remaining team. Evaporative cooling helps prevent some land-dwelling organisms from overheating. Evaporation from a plant's leaves keeps them from becoming too warm in the sun, just as sweating helps dissipate our excess body heat (**Figure 2.11**). On a much larger scale, the evaporation of surface waters cools tropical seas.

? **Explain the popular adage "It's not the heat, it's the humidity."**

● High humidity hampers cooling by slowing the evaporation of sweat.

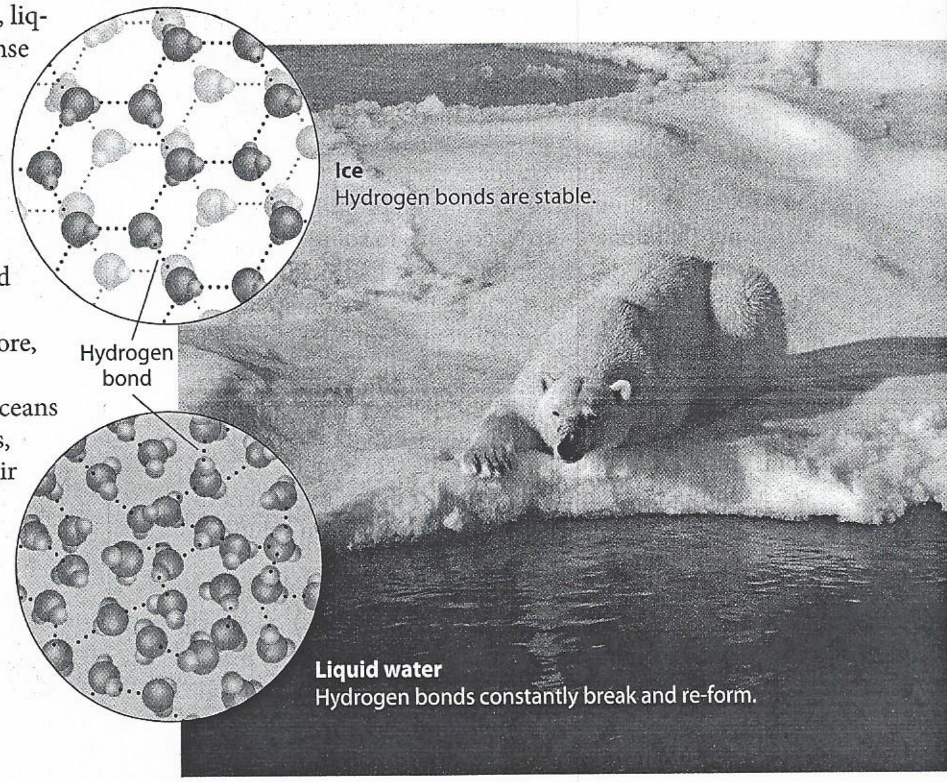
2.12 Ice floats because it is less dense than liquid water

Water exists on Earth in three forms: gas (water vapor), liquid, and solid. Unlike most substances, water is less dense as a solid than as a liquid. As you might guess, this unusual property is due to hydrogen bonds.

As water freezes, each molecule forms stable hydrogen bonds with its neighbors, holding them at "arm's length" and creating a three-dimensional crystal. In **Figure 2.12**, compare the spaciously arranged molecules in the ice crystal with the more tightly packed molecules in the liquid water. The ice crystal has fewer molecules than an equal volume of liquid water. Therefore, ice is less dense and floats on top of liquid water.

If ice sank, then eventually ponds, lakes, and even oceans would freeze solid. Instead, when a body of water cools, the floating ice insulates the water below from colder air above. This "blanket" of ice prevents the water from freezing and allows fish and other aquatic forms of life to survive under the frozen surface.

In the Arctic, this frozen surface serves as the winter hunting ground for polar bears (Figure 2.12). The shrinking of this ice cover as a result of global climate change may doom these bears.



▲ **Figure 2.12** Hydrogen bonds between water molecules in ice and water

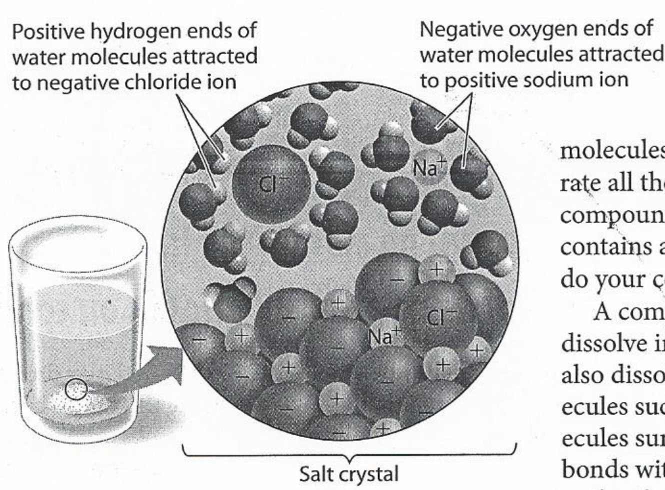
? Explain how freezing water can crack boulders.

Water in the crevices of a boulder expands as it freezes because the water molecules become spaced farther apart in forming ice crystals, which can crack the rock.

2.13 Water is the solvent of life

If you add a teaspoon of table salt to a glass of water, the salt will eventually dissolve, forming a solution. A **solution** is a liquid consisting of a uniform mixture of two or more substances. The dissolving agent (in our example, water) is the **solvent**, and a substance that is dissolved (in this case, salt) is a **solute**. An **aqueous solution** (from the Latin *aqua*, water) is one in which water is the solvent.

Water's versatility as a solvent results from the polarity of its molecules. **Figure 2.13** shows how a teaspoon of salt dissolves in water. At the surface of each grain, or crystal, the sodium and chloride ions are exposed to water. These ions and the water molecules are attracted to each other due to their opposite charges. The oxygen ends of the water molecules have a partial negative charge and cling to the positive sodium ions (Na). The hydrogen ends of the water molecules, with their partial positive charge, are attracted to the negative chloride ions (Cl). Working inward from the surface of each salt crystal, water molecules eventually surround and separate all the ions. Water dissolves other ionic compounds as well. Seawater, for instance, contains a great variety of dissolved ions, as do your cells.



▲ **Figure 2.13** A crystal of salt (NaCl) dissolving in water

A compound doesn't need to be ionic to dissolve in water. A spoonful of sugar will also dissolve in a glass of water. Polar molecules such as sugar dissolve as water molecules surround them and form hydrogen bonds with their polar regions. Even large molecules, such as proteins, can dissolve if they have ionic or polar regions on their surface. As the solvent inside all cells, in blood, and in plant sap, water dissolves an enormous variety of solutes necessary for life.

? Why are blood and most other biological fluids classified as aqueous solutions?

The solvent in these fluids is water.

2.14 The chemistry of life is sensitive to acidic and basic conditions

In liquid water, a small percentage of the water molecules dissociate or break apart into hydrogen ions (H^+) and hydroxide ions (OH^-). These ions are very reactive, and changes in their concentrations can drastically affect a cell's proteins and other complex molecules.

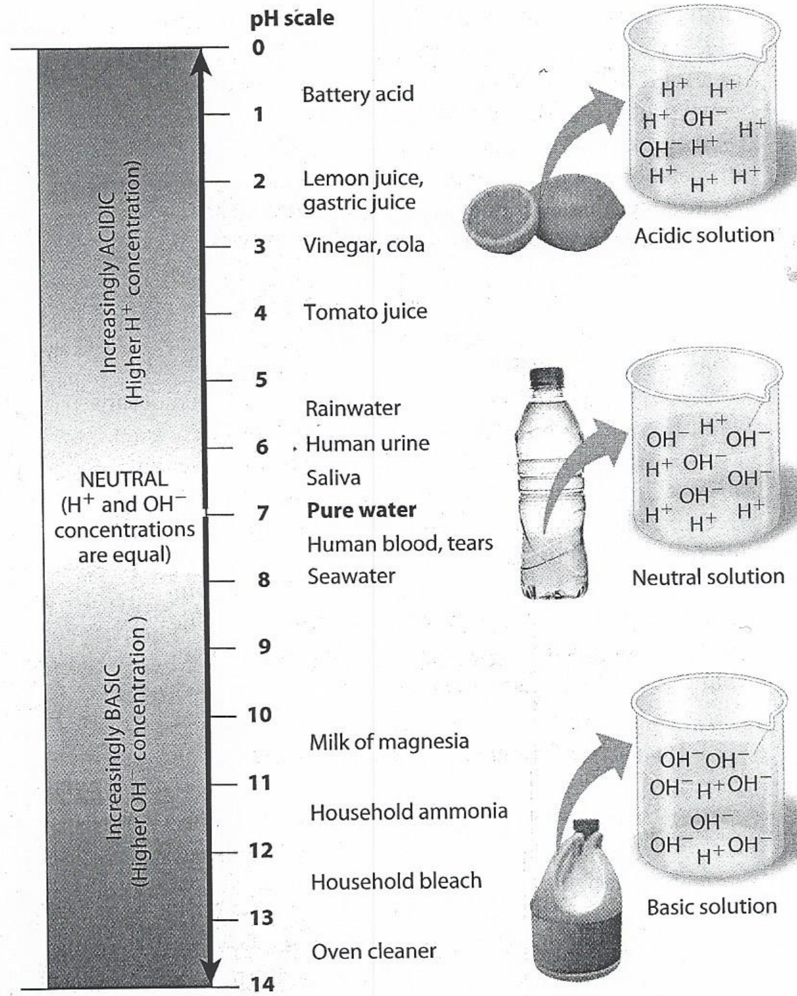
Some chemical compounds contribute additional H^+ to an aqueous solution, whereas others remove H^+ from it. A substance that donates hydrogen ions to solutions is called an **acid**. An example of a strong acid is hydrochloric acid (HCl), the acid in the gastric juice in your stomach. An acidic solution has a higher concentration of H^+ than OH^-.

A **base** is a substance that reduces the hydrogen ion concentration of a solution. Some bases, such as sodium hydroxide (NaOH), do this by donating OH^-; the OH^- combines with H^+ to form H_2O, thus reducing the H^+ concentration. Sodium hydroxide is a common ingredient in oven cleaners. Other bases accept H^+ ions from solution, resulting in a higher OH^- concentration.

We use the **pH scale** to describe how acidic or basic a solution is (pH stands for potential of hydrogen). As shown in **Figure 2.14**, the scale ranges from 0 (most acidic) to 14 (most basic). Each pH unit represents a 10-fold change in the concentration of H^+ in a solution. For example, lemon juice at pH 2 has 10 times more H^+ than an equal amount of a cola at pH 3 and 100 times more H^+ than tomato juice at pH 4.

Pure water and aqueous solutions that are neither acidic nor basic are said to be neutral; they have a pH of 7, and the concentrations of H^+ and OH^- are equal. The pH inside most cells is close to 7.

The pH of human blood plasma (the fluid portion of the blood) is very close to 7.4. A person cannot survive for more than a few minutes if the blood pH drops to 7.0 or rises to 7.8. How can your body maintain a relatively constant pH in your cells and blood? Biological fluids contain **buffers**, substances that minimize changes in pH. They do so by accepting H^+ when it is in excess and donating H^+ when it is depleted.



▲ **Figure 2.14** The pH scale, which reflects the relative concentrations of H^+ and OH^-

? **Compared to a basic solution at pH 9, the same volume of an acidic solution at pH 4 has ____________ times more H^+.**

100,000

2.15 Scientists study the effects of rising atmospheric CO_2 on coral reef ecosystems

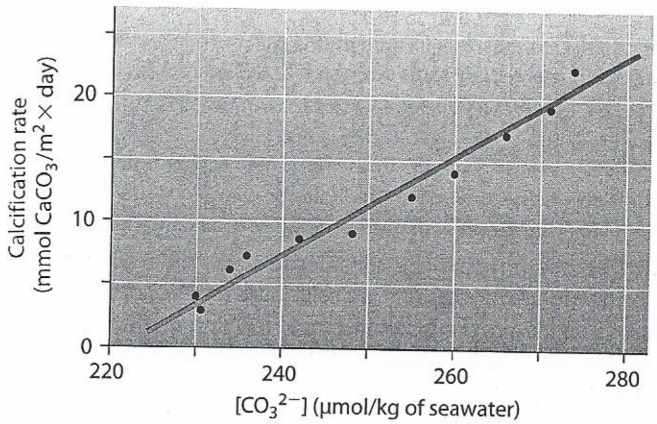
SCIENTIFIC THINKING

Will rising atmospheric CO_2 harm coral reefs?

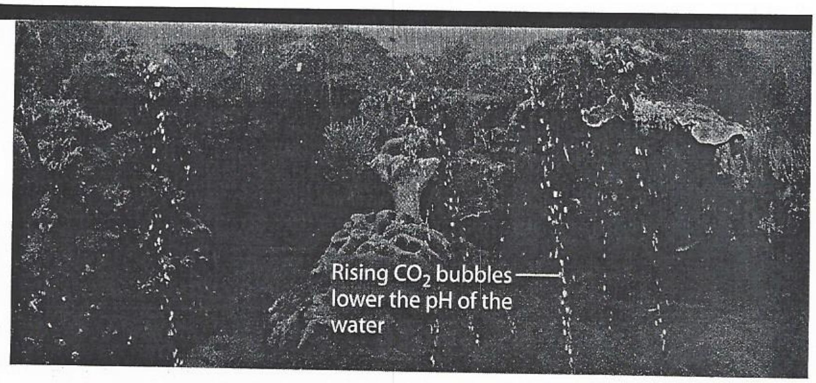
Carbon dioxide is the main product of fossil fuel combustion, and its steadily increasing release into the atmosphere is linked to global climate change (see Module 38.4). About 25% of this CO_2 is absorbed by the oceans—and this naturally occurring remedy to excess CO_2 would seem to be a good thing. However, as CO_2 levels on the planet continue to rise, the increasing absorption of CO_2 is expected to change ocean chemistry and harm marine life and ecosystems.

In **ocean acidification**, CO_2 dissolving in seawater lowers the pH of the ocean. Recent studies estimate that the pH of the oceans is 0.1 pH unit lower now than at any time in the past 420,000 years and may drop from the current level of 8.1 to 7.8 by the end of this century. How will this affect marine organisms?

Several studies investigating the impact of a lower pH on coral reef ecosystems have looked at the process called calcification, in which coral animals combine calcium and carbonate ions to form their calcium carbonate skeletons. As seawater acidifies, the extra hydrogen ions (H^+) combine with carbonate ions (CO_3^{2-}) to form bicarbonate ions (HCO_3^-). This reaction reduces the carbonate ion concentration available to corals and other shell-building animals. Scientists predict that ocean acidification will cause the carbonate ion concentration to decrease by 40% by the year 2100.



▲ **Figure 2.15A** The effect of carbonate ion concentration on calcification rate in an artificial coral reef system. (The independent variable shown on the *x* axis is the concentration of carbonate ions, which the researchers manipulated. The dependent variable—the calcification rate, shown on the *y* axis—is what was measured in the experiment and was predicted to "depend on" or respond to the experimental treatment.)



▲ **Figure 2.15B** A "champagne" reef with bubbles of CO_2 rising from a volcanic seep

In a controlled experiment, scientists looked at the effect of decreasing carbonate ion concentration on the rate of calcium deposition by reef organisms. The Biosphere 2 aquarium in Arizona contains a large coral reef system that behaves like a natural reef. Researchers measured how the calcification rate changed with differing amounts of dissolved carbonate ions. **Figure 2.15A** presents the results of one set of experiments, in which pH, temperature, and concentration of calcium ions were held constant while the carbonate ion concentration of the seawater was varied. As you can see from the graph, the lower the concentration of carbonate ions, the lower the rate of calcification, and the slower the growth of coral animals.

Controlled studies such as this one have provided evidence that ocean acidification and the resulting reduction in carbonate ion concentration will negatively affect coral reefs. But scientists have also looked to natural habitats to study how ocean acidification affects coral reef ecosystems. A 2011 study looked at three volcanic seeps in Papua New Guinea. As you can see in **Figure 2.15B**, bubbles of CO_2 released from underwater volcanoes around such "champagne reefs" lower the pH of the water. Researchers surveyed three study sites in which the pH naturally varied from 8.1 to 7.8. They found reductions in coral diversity and the recruitment of juvenile coral as the pH of the sites declined, both of which undermine the resiliency of a reef community. Researchers also found a shift to less structurally complex and slower growing corals. The structural complexity of coral reef ecosystems makes them havens for a great diversity of organisms.

Scientists often synthesize their conclusions using multiple lines of evidence. The results from both controlled experimental studies and observational field studies of sites where pH naturally varies have dire implications for the health of coral reefs and the diversity of organisms they support.

? **What is the relationship between fossil fuel consumption and coral reefs?**

● Some of the increased CO_2 released by burning fossil fuels dissolves in and lowers the pH of the oceans. A lower pH reduces levels of carbonate ions, which then lowers the rate of calcification by coral animals. A lower pH also changes the composition and resiliency of coral reefs.

2.16 The search for extraterrestrial life centers on the search for water

EVOLUTION CONNECTION

When astrobiologists search for signs of extraterrestrial life on distant planets, they look for evidence of water. Why? As we've seen in this chapter, the emergent properties of water support life on Earth in many ways. Is it possible that some form of life has evolved on other planets that have water in their environment? Scientists with the National Aeronautics and Space Administration (NASA) are looking into this possibility.

Like Earth, Mars has an ice cap at both poles, and scientists have found signs that water may exist elsewhere on the planet. In 2008, the robotic spacecraft *Phoenix* landed on Mars and sent back images showing that ice is present just under Mars's surface. Then, in 2011, high-resolution images sent to Earth from the Mars Reconnaissance Orbiter showed evidence for liquid water beneath the surface. Distinctive streaks form along steep slopes during the Mars spring and summer, which then vanish during the winter. Careful study of these images over time has led scientists to conclude that these streaks are most likely seasonal streams of flowing water resulting when subsurface ice melts during the warm season. This exciting finding has reinvigorated the search for signs of life, past or present, on Mars and other planets. If any life-forms or fossils are found, their study will shed light on the process of evolution from an entirely new perspective.

? **Why is the presence of water important in the search for extraterrestrial life?**

● Water plays important roles in life as we know it, from moderating temperatures on the planet to functioning as the solvent of life.

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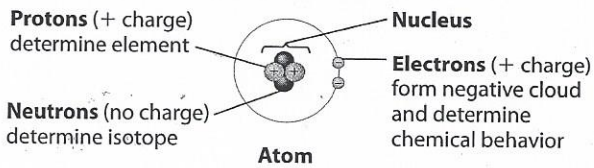
Reviewing the Concepts

Elements, Atoms, and Compounds (2.1–2.4)

2.1 Organisms are composed of elements, in combinations called compounds. Oxygen, carbon, hydrogen, and nitrogen make up about 96% of living matter.

2.2 Trace elements are common additives to food and water.

2.3 Atoms consist of protons, neutrons, and electrons.



2.4 Radioactive isotopes can help or harm us. Radioactive isotopes are valuable in basic research and medicine.

Chemical Bonds (2.5–2.9)

2.5 The distribution of electrons determines an atom's chemical properties. An atom whose outer electron shell is not full tends to interact with other atoms and share, gain, or lose electrons, resulting in attractions called chemical bonds.

2.6 Covalent bonds join atoms into molecules through electron sharing. In a nonpolar covalent bond, electrons are shared equally. In polar covalent bonds, such as those found in water, electrons are pulled closer to the more electronegative atom.

2.7 Ionic bonds are attractions between ions of opposite charge. Electron gain and loss create charged atoms, called ions.

2.8 Hydrogen bonds are weak bonds important in the chemistry of life. The slightly positively charged H atoms in one polar molecule may be attracted to the partial negative charge of an oxygen or nitrogen atom in a neighboring molecule.

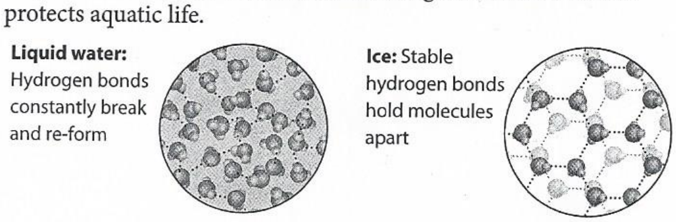
2.9 Chemical reactions make and break chemical bonds. The composition of matter is changed as bonds are broken and formed to convert reactants to products.

Water's Life-Supporting Properties (2.10–2.16)

2.10 Hydrogen bonds make liquid water cohesive. Cohesion creates surface tension and allows water to move from plant roots to leaves.

2.11 Water's hydrogen bonds moderate temperature. Heat is absorbed when hydrogen bonds break and released when hydrogen bonds form. This helps keep temperatures relatively steady. As the most energetic water molecules evaporate, the surface of a substance cools.

2.12 Ice floats because it is less dense than liquid water. Floating ice protects lakes and oceans from freezing solid, which in turn protects aquatic life.



2.13 Water is the solvent of life. Polar or charged solutes dissolve when water molecules surround them, forming aqueous solutions.

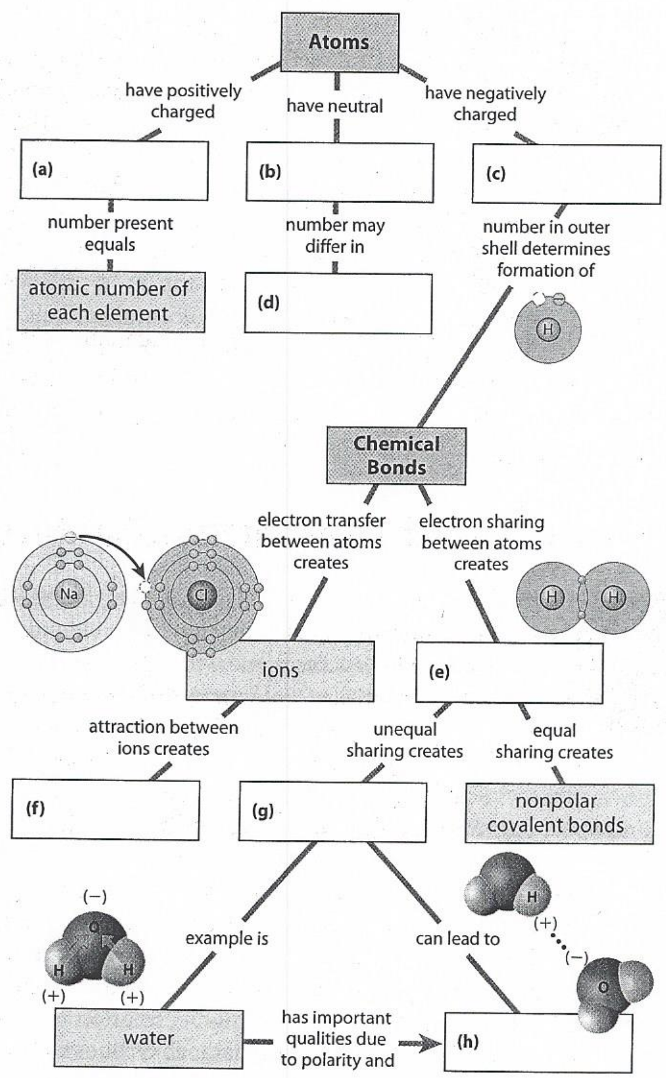
2.14 The chemistry of life is sensitive to acidic and basic conditions. A compound that releases H^+ in solution is an acid, and one that accepts H^+ is a base. The pH scale ranges from 0 (most acidic) to 14 (most basic). The pH of most cells is close to 7 (neutral) and is kept that way by buffers.

2.15 Scientists study the effects of rising atmospheric CO_2 on coral reef ecosystems. The acidification of the ocean threatens coral reefs and other marine organisms.

2.16 The search for extraterrestrial life centers on the search for water. The emergent properties of water support life on Earth and may contribute to the potential for life to have evolved on other planets.

Connecting the Concepts

1. Fill in the blanks in this concept map to help you tie together the key concepts concerning elements, atoms, and molecules.



2. Create a concept map to organize your understanding of the life-supporting properties of water. A sample map is in the answer section, but the value of this exercise is in the thinking and integrating you must do to create your own map.

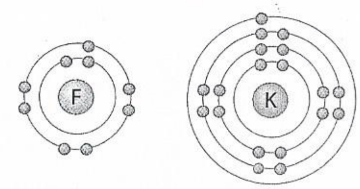
Testing Your Knowledge

Level 1: Knowledge/Comprehension

3. Changing the ______________ would change it into an atom of a different element.
   a. number of electrons surrounding the nucleus of an atom
   b. number of protons in the nucleus of an atom
   c. electrical charge of an atom
   d. number of neutrons in the nucleus of an atom
4. A solution at pH 6 contains ______________ H^+ than the same amount of a solution at pH 8.
   a. 20 times more
   b. 100 times more
   c. 2 times less
   d. 100 times less
5. Most of the unique properties of water result from the fact that water molecules
   a. are the most abundant molecules on Earth's surface.
   b. are held together by covalent bonds.
   c. are constantly in motion.
   d. are polar and form hydrogen bonds.
6. A can of cola consists mostly of sugar dissolved in water, with some carbon dioxide gas that makes it fizzy and makes the pH less than 7. In chemical terms, you could say that cola is an aqueous solution where water is the ______________, sugar is a ______________, and carbon dioxide makes the solution ______________.
   a. solvent . . . solute . . . basic
   b. solute . . . solvent . . . basic
   c. solvent . . . solute . . . acidic
   d. solute . . . solvent . . . acidic

Level 2: Application/Analysis

7. The atomic number of sulfur (S) is 16. Sulfur combines with hydrogen by covalent bonding to form a compound, hydrogen sulfide. Based on the number of valence electrons in a sulfur atom, predict the molecular formula of the compound. (*Explain your answer.*)
   a. HS
   b. H_2S
   c. H_4S_2
   d. H_4S
8. In what way does the need for iodine or iron in your diet differ from your need for calcium or phosphorus?
9. Use carbon-12, the most common isotope of carbon, to define these terms: atomic number, mass number, valence. Which of these numbers is most related to the chemical behavior of an atom? Explain.
10. In terms of electron sharing between atoms, compare nonpolar covalent bonds, polar covalent bonds, and ions.
11. The diagram below shows the arrangement of electrons around the nucleus of a fluorine and a potassium atom. What kind of bond do you think would form between these two atoms?



Level 3: Synthesis/Evaluation

12. Look back at the abbreviated periodic table of the elements in Figure 2.5B. If two elements are in the same row, what do they have in common? If two elements are in the same column, what do they have in common? Would you predict that elements in the same row or the same column will have similar chemical properties? Explain.
13. **SCIENTIFIC THINKING** A recent experimental study looked at the combined effects of ocean acidification (see Module 2.15) and increased ocean temperatures, both aspects of global climate change, on the growth of polyps, juvenile coral animals. Researchers reported the average polyp biomass (in μg/polyp) after 42 days of growth under four treatments: a control with pH and temperature maintained close to normal reef conditions, a pH lowered by 0.2 units, a temperature raised by 1°C, and a combined lower pH and higher temperature. The results showed that polyp biomass was reduced somewhat in both the low-pH and high-temperature treatments, but the combined treatment resulted in a reduction in growth by almost a third—a statistically significant result. Experiments often look at the effects of changing one variable at a time, while keeping all other variables constant. Explain why this experiment considered two variables—both a higher temperature and a lower pH—at the same time.
14. In agricultural areas, farmers pay close attention to the weather forecast. Right before a predicted overnight freeze, farmers spray water on crops to protect the plants. Use the properties of water to explain how this method works. Be sure to mention why hydrogen bonds are responsible for this phenomenon.
15. This chapter explains how the emergent properties of water contribute to the suitability of the environment for life. Until fairly recently, scientists assumed that other physical requirements for life included a moderate range of temperature, pH, and atmospheric pressure. That view has changed with the discovery of organisms known as extremophiles, which have been found flourishing in hot, acidic sulfur springs and around hydrothermal vents deep in the ocean. What does the existence of life in such environments say about the possibility of life on other planets?

Answers to all questions can be found in Appendix 4.