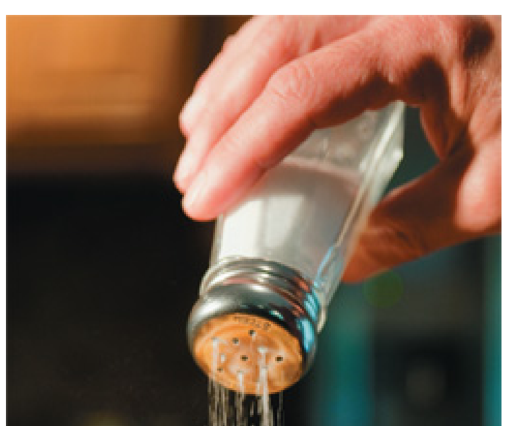
2
Essential Chemistry for Biology

Why Chemistry Matters

Too little of the essential element copper in your diet causes anemia, but too much causes kidney and liver damage.

Sodium is an explosive solid and chlorine is a poisonous gas, but when combined they form a common ingredient in your diet: table salt.



Lemon juice has roughly the same acidity as the food-digesting chemicals in your stomach.

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Radioactivity: Biology and Society

Radiation and Health

The word "radioactive" probably sets off alarm bells in your mind: "Danger! Hazardous!" It is true that radiation, high-energy particles emitted by radioactive substances, can penetrate living tissues and kill cells by damaging DNA. But you probably know that radiation can also be medically beneficial by, for example, helping to treat cancer. So what determines whether radiation is harmful or helpful to an organism's health?

Radiation is most dangerous when exposure is uncontrolled and covers most or all of the body, as happens when a person is exposed to radioactive fallout from a nuclear detonation or accident. In contrast, controlled medical radiation therapy exposes only a small part of the body to a precise dosage of radiation. For example, when treating cancer, carefully calibrated radiation beams are aimed from several angles, intersecting only at the tumor. This provides a deadly dose to cancerous cells but mostly spares surrounding healthy tissues. Radiation therapy is also used to treat Graves' disease, a condition in which an overactive thyroid gland (located in the neck) causes a variety of physical symptoms, including shaking, swelling behind the eyeballs, and heart irregularities. Individuals with Graves' disease may be treated with a "cocktail" containing radioactive iodine.

Because the thyroid produces a variety of hormones that use iodine, the radioactive iodine accumulates in this gland, where it then provides a steady low dose of radiation that can, over time, destroy enough thyroid tissue to reduce symptoms.

Pros and cons of radiation.

Radiation can be harmful if it is released into the environment, but physicians use controlled doses of radiation to diagnose and treat several diseases.

What makes something radioactive? To understand this question, we have to look to the most basic level of all living things: the atoms that make up all matter. Many questions about life—for example, why is radiation harmful to cells?—can be reduced to questions about chemicals and their interactions—for example, how does radiation affect the atoms in living issues? Knowledge of chemistry is therefore essential to understanding biology. In this chapter, we'll review some basic chemistry that you can apply throughout your study of life. We'll start with an examination of molecules, atoms, and their components. Next, we'll discuss water, one of life's most important molecules, and its crucial role in sustaining life on Earth.

Some Basic Chemistry

Why would a biology textbook include a chapter on chemistry? Well, take any biological system apart and you eventually end up at the chemical level. In fact, you can think of your body as a big watery container of chemicals undergoing a continuous series of chemical reactions. Viewed this way, your metabolism—the sum total of all the chemical reactions that occur in your body—is like a giant square dance, with chemical partners constantly swapping atoms as they move to and fro. Beginning at this basic biological level, let's explore the chemistry of life.

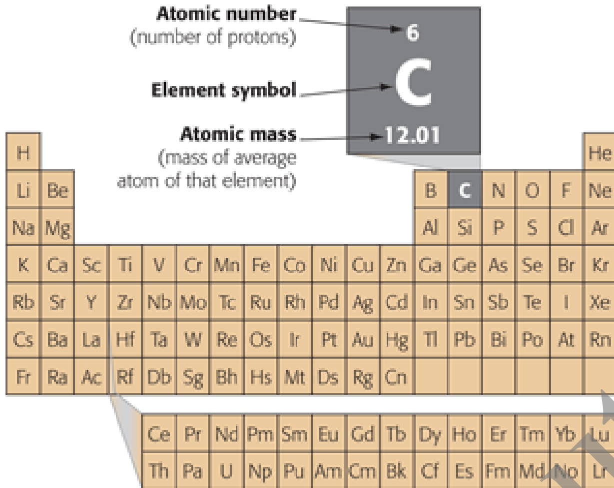
Matter: Elements and Compounds

You and everything that surrounds you is made of matter, the physical "stuff" of the universe. Matter is found on Earth in three physical states: solid, liquid, and gas. Defined more formally, **matter** is anything that occupies space and has mass. **Mass** is a measure of the amount of material in an object. All matter is composed of chemical elements. An **element** is a substance that cannot be broken down into other substances by chemical reactions. Think of it this way: When you burn wood, you are left with ash. But when you burn ash, you only get more ash. That is because wood is a complex mixture of elements, while ash is a pure element (carbon) that cannot be further broken down. There are 92 naturally occurring elements; examples are carbon, oxygen, and gold. Each element has a symbol derived from its English, Latin, or German name. For instance, the symbol for gold, Au, is from the Latin word *aurum*. All the elements —the 92 that occur naturally and several dozen that are human-made—are listed in the **periodic table of the elements**, a familiar fixture in any chemistry or biology lab (Figure 2.1; see Appendix B for a full version).

Too little of the essential element copper in your diet causes anemia, but too much causes kidney and liver damage.

Figure 2.1 **Abbreviated periodic table of the elements.**

In the full periodic table (see Appendix B), each entry contains the element symbol in the center, with the atomic number above and the atomic mass below. The element highlighted here is carbon (C).



Mercury (Hg)

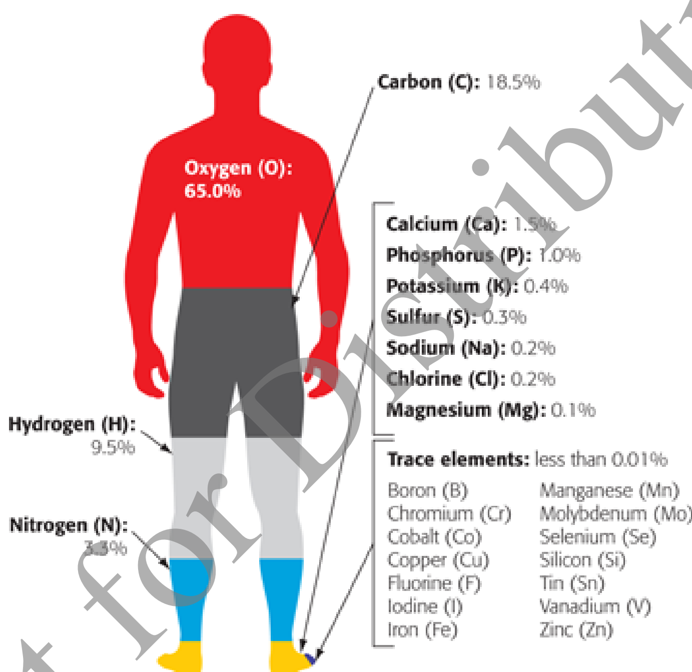
Copper (Cu)

Lead (Pb)

Of the naturally occurring elements, 25 are essential to people. (Other organisms need fewer; plants, for example, typically need 17.) Four of these elements—oxygen (O), carbon (C), hydrogen (H), and nitrogen (N)—make up about 96% of the weight of the body (Figure 2.2). Much of the remaining 4% is accounted for by 7 elements, most of which are probably familiar to you, such as calcium (Ca). Calcium, important for building strong bones and teeth, is found abundantly in milk and dairy products as well as sardines and green, leafy vegetables (kale and broccoli, for example).

Figure 2.2 Chemical composition of the human body.

Notice that just 4 elements make up 96% of your weight.



Less than 0.01% of your weight is made up of 14 trace elements. **Trace elements** are required in only very small amounts, but you cannot live without them. The average person, for example, needs only a tiny speck of iodine each day. Iodine is an essential ingredient of hormones produced by the thyroid gland, located in the neck. An iodine deficiency causes the thyroid gland to enlarge, a condition called goiter. Therefore, consuming foods that are naturally rich in iodine—such as green vegetables, eggs, kelp, and dairy products—prevents goiter. The addition of iodine to table salt ("iodized salt") has nearly eliminated goiter in industrialized nations, but many thousands of people in developing countries are still affected (Figure 2.3). Another trace element is fluorine, which (in the form of fluoride) is added to dental products and drinking water and helps maintain healthy bones and teeth. Many prepared foods are fortified with trace mineral elements. Look at the side of a cereal box and you'll probably see iron listed; you can actually see the iron yourself if you crush the cereal and stir a magnet through it. Be grateful for this additive: It helps prevent anemia due to iron deficiency, one of the most common nutritional deficiencies among Americans.

Figure 2.3 Diet and goiter.

Goiter, an enlargement of the thyroid gland, shown here in a Malaysian woman, can occur when a person's diet does not include enough iodine, a trace element.

Eating iodine-rich foods can prevent goiter.

Elements can combine to form **compounds**, substances that contain two or more elements in a fixed ratio. In everyday life, compounds are much more common than pure elements. Familiar examples of compounds include table salt and water. Table salt is sodium chloride, NaCl, consisting of equal parts of the elements sodium (Na) and chlorine (Cl). A molecule of water, H_2O, has two atoms of hydrogen and one atom of oxygen. Most of the compounds in living organisms contain several different elements. DNA, for example, contains carbon, nitrogen, oxygen, hydrogen, and phosphorus.

Sodium is an explosive solid and chlorine is a poisonous gas, but when combined they form a common ingredient in your diet: table salt.

Checkpoint

How many of the naturally occurring elements are used by your body? Which four are the most abundant in living cells?

Answer: *25; oxygen, carbon, hydrogen, and nitrogen*

Atoms

Each element is made up of one kind of atom, and the atoms in an element are different from the atoms of other elements. An **atom** is the smallest unit of matter that still retains the properties of an element. In other words, the smallest amount of the element carbon is one carbon atom. Just how small is this "piece" of carbon? It would take about a million carbon atoms to stretch across the period at the end of this sentence.

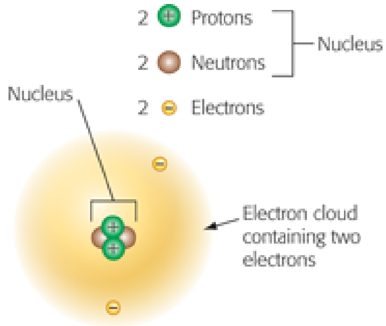
The Structure of Atoms

Atoms are composed of subatomic particles, of which the three most important are protons, electrons, and neutrons. A **proton** is a subatomic particle with a single unit of positive electrical charge (+). An **electron** is a subatomic particle with a single negative charge (–). A **neutron** is electrically neutral (has no charge).

Figure 2.4 shows a simplified model of an atom of the element helium (He), the lighter-than-air gas used to inflate party balloons. Each atom of helium has 2 neutrons (●) and 2 protons (⊕) tightly packed into the **nucleus**, the atom's central core. Two electrons (⊖) move around the nucleus in a spherical cloud at nearly the speed of light. The electron cloud is much bigger than the nucleus. If the atom were the size of a baseball stadium, the nucleus would be the size of a pea on the pitcher's mound and the electrons would be two gnats buzzing around the bleachers. When an atom has an equal number of protons and electrons, its net electrical charge is zero and so the atom is neutral.

Figure 2.4 A simplified model of a helium atom.

This model shows the subatomic particles in an atom of helium. The electrons move very fast, creating a spherical cloud of negative charge surrounding the positively charged nucleus.



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, and no other element has 2 protons. The periodic table of elements (Appendix B) lists elements in order of atomic number. Note that in these atoms, the atomic number is also the number of electrons. A standard atom of any element has an equal number of protons and electrons, and thus its net electrical charge is 0 (zero). An atom's **mass number** is the sum of the number of protons and neutrons. 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 its mass is approximated as zero. An atom's **atomic mass**, which is listed in the periodic table as the bottom number (under the element symbol), is close to its mass number—the sum of its protons and neutrons—but may differ slightly because it represents an average of all the naturally occurring forms of that element.

Isotopes

Some elements can exist in different forms called **isotopes**, which have the same numbers of protons and electrons as a standard atom of that element but different numbers of neutrons. In other words, isotopes are forms of an element that differ in mass. As shown in Table 2.1, the isotope carbon-12 (named for its mass number), which has 6 neutrons and 6 protons, makes up about 99% of all naturally occurring carbon. Most of the other 1% of carbon on Earth is the isotope carbon-13, which has 7 neutrons and 6 protons. A third isotope, carbon-14, which has 8 neutrons and 6 protons, occurs in minute quantities. 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 radioactive. A **radioactive isotope** is one in which the nucleus decays spontaneously, shedding particles and energy.

Table 2.1 **Isotopes of Carbon**

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

Radiation from decaying isotopes can damage cellular molecules and thus can pose serious health risks. In 1986, the explosion of a nuclear reactor at Chernobyl, Ukraine, released large amounts of radioactive isotopes, killing 30 people within a few weeks. Millions of people in the surrounding areas were exposed, causing an estimated 6,000 cases of thyroid cancer. The 2011 post-tsunami Fukushima nuclear disaster in Japan, did not result in any immediate deaths due to radiation exposure, but scientists are carefully monitoring the people who live in the area to watch for any long-term health consequences.

Natural sources of radiation can also pose a threat. Radon, a radioactive gas, can cause lung cancer. Radon may contaminate buildings where underlying rocks naturally contain the radioactive element uranium. Homeowners can install a radon detector or test their home to ensure that radon levels are safe.

Although radioactive isotopes can cause harm when uncontrolled, they have many uses in biological research and medicine. In the Biology and Society section, we discussed how radioactivity can be used to treat diseases such as cancer and Graves' disease. Let's take a look at another beneficial use of radioactivity: the diagnosis of disease.

Checkpoint

By definition, all atoms of carbon have exactly 6 ______ , but the number of ______ varies from one isotope to another.

Answer: *protons; neutrons*

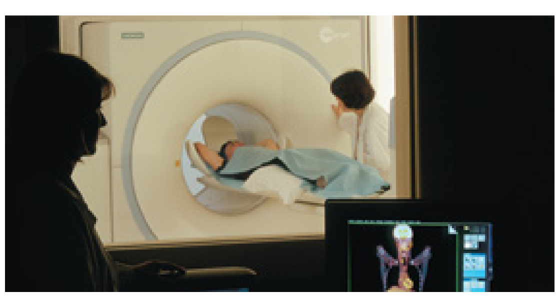
Radioactivity: The Process of Science

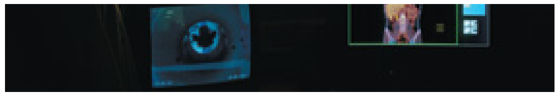
Can Radioactive Tracers Identify Brain Diseases?

Cells use radioactive isotopes the same way they use nonradioactive isotopes of the same element. Once the cell takes up a radioactive isotope, the location and concentration of the isotope can be detected because of the radiation it emits. This makes radioactive isotopes useful as tracers—biological spies, in effect—for monitoring living organisms. For example, a medical diagnostic tool called a PET scan works by detecting small amounts of radiation emitted by radioactive materials that were purposefully introduced into the body (Figure 2.5).

Figure 2.5 A PET scan.

This monitor shows the images produced by a PET scanner. PET scans can be used to diagnose several diseases, including epilepsy, cancer, and Alzheimer's disease.





In 2012, researchers published a study that used PET scans to investigate Alzheimer's disease. In Alzheimer's disease, a patient gradually loses his or her memory and can become confused, forgetful, and unable to perform normal daily tasks. The disease inevitably leads to a loss of bodily functions and death. A definitive diagnosis of Alzheimer's is difficult because it is hard to distinguish from other age-related disorders. Early detection and treatment of Alzheimer's could benefit many patients and their families.

The **observation** that the brains of people with Alzheimer's are often filled with clumps of a protein called amyloid led the researchers to **question** whether these clumps could be detected by a PET scan. The researchers formed the **hypothesis** that a molecule called florbetapir, which contains the radioactive isotope fluorine-18, could be detected by PET scans after it binds to amyloid deposits in living patients. The researchers' **prediction** was that using florbetapir during PET scans could help with diagnosis.

Their **experiment** involved 229 patients who had been diagnosed with mental decline. Of these, 113 patients showed amyloid deposits in their PET scans. This information led doctors to change the diagnosis in 55% of the patients, sometimes changing the diagnosis to Alzheimer's and sometimes changing it to a different disease. Furthermore, the PET scan data led to changes in treatment (such as different drugs) in 87% of cases. These **results** indicate that radioisotope scans can indeed alter diagnoses and affect treatment. Researchers hope that this will lead to improved outcomes for patients suffering from this debilitating condition.

Chemical Bonding and Molecules

Of the three subatomic particles we've discussed—protons, neutrons, and

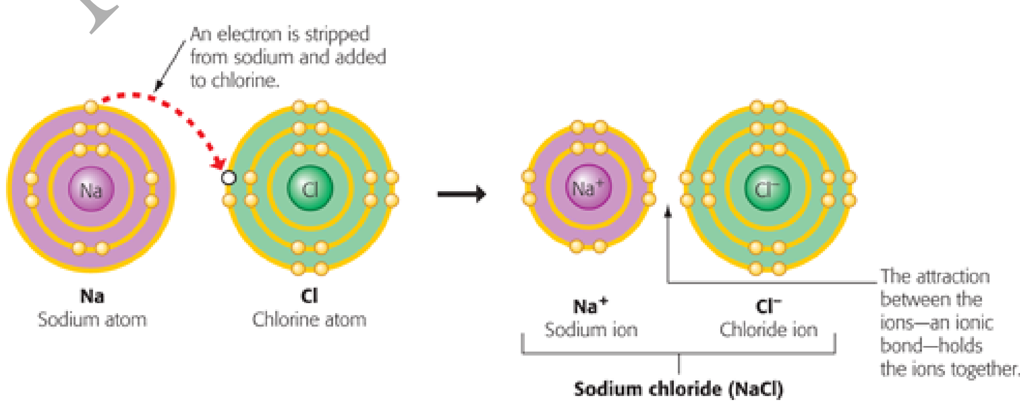
Of the three subatomic particles we've discussed—protons, neutrons, and electrons—only electrons are directly involved in chemical reactions. The number of electrons in an atom determines the chemical properties of that atom. Chemical reactions enable atoms to transfer or share electrons. These interactions usually result in atoms staying close together, held by attractions called **chemical bonds**. In this section, we will discuss three types of chemical bonds: ionic, covalent, and hydrogen bonds.

Ionic Bonds

Table salt is an example of how the transfer of electrons can bond atoms together. As discussed earlier, the two ingredients of table salt are the elements sodium (Na) and chlorine (Cl). When in close proximity to each other, a chlorine atom strips an electron from a sodium atom (Figure 2.6). Before this electron transfer, both the sodium and chlorine atoms are electrically neutral. Because electrons are negatively charged, the electron transfer moves one unit of negative charge from sodium to chlorine. This action makes both atoms **ions**, atoms or molecules that are electrically charged as a result of gaining or losing electrons. In this case, the loss of an electron results in the sodium ion having a charge of +1, whereas chlorine's gain of an electron results in it having a charge of –1. The sodium ion (Na^+) and chloride ion (Cl^-) are then held together by an **ionic bond**, the attraction between oppositely charged ions. Compounds, such as table salt, that are held together by ionic bonds are called ionic compounds. (Note that negatively charged ions often have names ending in "-ide," like "chloride" or "fluoride.")

Figure 2.6 Electron transfer and ionic bonding.

When a sodium atom and a chlorine atom meet, the electron transfer between the two atoms results in two ions with opposite charges.



Checkpoint

When a lithium ion (Li^+) joins a bromide ion (Br^-) to form lithium bromide, the resulting bond is a(n) ______ bond.

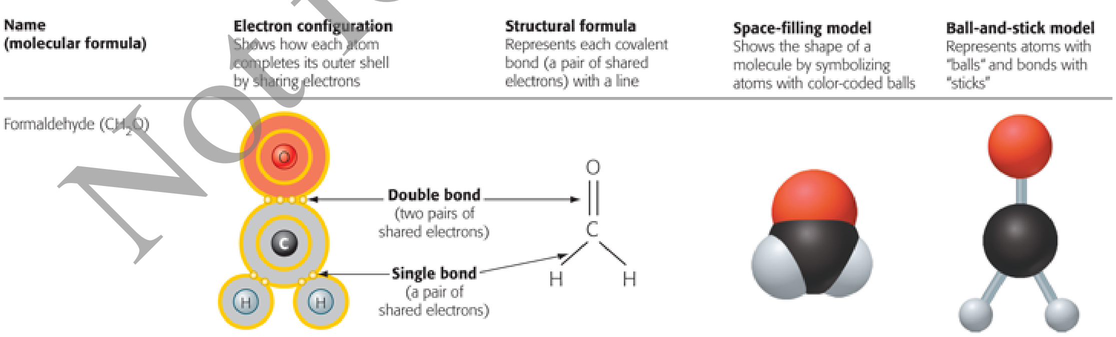
Answer: *ionic*

Covalent Bonds

In contrast to the complete *transfer* of electrons in ionic bonds, a **covalent bond** forms when two atoms *share* one or more pairs of electrons. Of the bonds we've discussed, covalent bonds are the strongest; these are the bonds that hold atoms together in a **molecule**. For example, in Figure 2.7, you can see that each of the two hydrogen atoms in a molecule of formaldehyde (CH_2O, a common disinfectant and preservative) shares one pair of electrons with the carbon atom. The oxygen atom shares two pairs of electrons with the carbon, forming a double bond. Notice that each atom of hydrogen (H) can form one covalent bond, oxygen (O) can form two, and carbon (C) can form four.

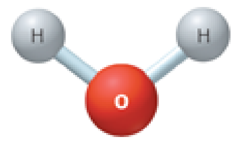
Figure 2.7 Alternative ways to represent a molecule.

A molecular formula, such as CH_2O, tells you the number of each kind of atom in a molecule but not how they are attached together. This figure shows four common ways of representing the arrangement of atoms in molecules.

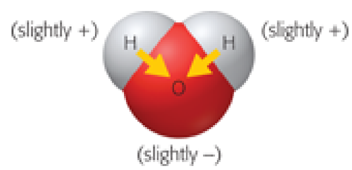


Hydrogen Bonds

A molecule of water (H_2O) consists of two hydrogen atoms joined to one oxygen atom by single covalent bonds. (The covalent bonds between the atoms are represented as the "sticks" here in a ball-and-stick illustration, and the atoms are shown as the "balls"):



However, the electrons are not shared equally between the oxygen and hydrogen atoms. The two yellow arrows shown in the space-filling model here indicate the stronger pull on the shared electrons that oxygen has compared with its hydrogen partners:

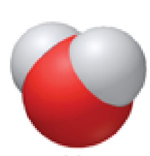


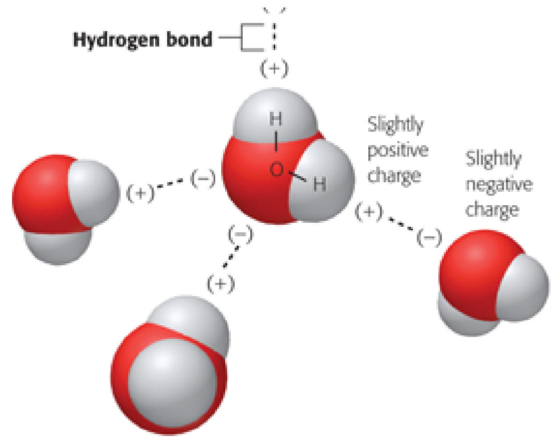
The unequal sharing of negatively charged electrons, combined with its V shape, makes a water molecule polar. A **polar molecule** is one with an uneven distribution of charge that creates two poles, one positive pole and one negative pole. In the case of water, the oxygen end of the molecule has a slight negative charge, and the region around the two hydrogen atoms is slightly positive.

The polarity of water results in weak electrical attractions between neighboring water molecules. Because opposite charges attract, water molecules tend to orient such that a hydrogen atom from one water molecule is near the oxygen atom of an adjacent water molecule. These weak attractions are called **hydrogen bonds** (Figure 2.8). As you will see later in this chapter, the ability of water to form hydrogen bonds has many crucial implications for life on Earth.

Figure 2.8 Hydrogen bonding in water.

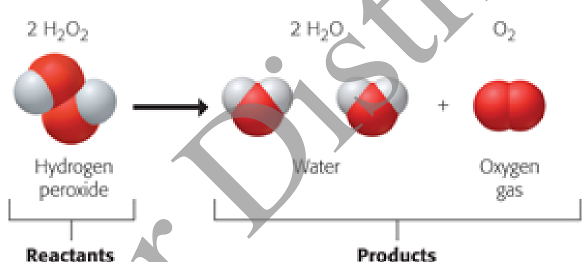
The charged regions of the polar water molecules are attracted to oppositely charged areas of neighboring molecules. Each molecule can hydrogen-bond to a maximum of four partners.





Chemical Reactions

The chemistry of life is dynamic. Your cells are constantly rearranging molecules by breaking existing chemical bonds and forming new ones in a "chemical square dance." Such changes in the chemical composition of matter are called **chemical reactions**. An example of a chemical reaction is the breakdown of hydrogen peroxide (a common disinfectant that you may have poured on a cut):



Let's translate the chemical shorthand: Two molecules of hydrogen peroxide (2 H_2O_2) react to form two molecules of water (2 H_2O) and one molecule of oxygen (O_2, which is responsible for the fizzing that happens when hydrogen peroxide interacts with blood). The arrow in this equation indicates the conversion of the starting materials, the **reactants** (2 H_2O_2), to the **products** (2 H_2O and O_2).

Notice that the same total numbers of hydrogen and oxygen atoms are present in reactants (to the left of the arrow) and products (to the right), although they are grouped differently. Chemical reactions cannot create or destroy matter; they can only rearrange it. These rearrangements usually involve the breaking of chemical bonds in reactants and the forming of new bonds in products.

This discussion of water molecules as the product of a chemical reaction is a good conclusion to this section on basic chemistry. Water is a substance so important in biology that we'll take a closer look at its life-supporting properties in the next section.

Checkpoint

Predict the formula for the compound that results when a molecule of sulfur trioxide (SO_3) combines with a molecule of water to produce a single molecule of product. (*Hint*: In chemical reactions, no atoms are gained or lost.)

Answer: *H_2SO_4 (sulfuric acid, which is a component of acid rain)*

Water and Life

Life on Earth began in water and evolved there for 3 billion years before spreading onto land. Modern life, even land-dwelling life, is still tied to water. You've had personal experience with this dependence on water every time you seek liquids to quench your thirst. Inside your body, your cells are surrounded by a fluid that's composed mostly of water, and your cells themselves range from 70% to 95% in water content.

The abundance of water is a major reason that Earth is habitable. Water is so common that it is easy to overlook the fact that it is an exceptional substance with many extraordinary properties (Figure 2.9). We can trace water's unique life-supporting properties to the structure and interactions of its molecules.

Figure 2.9 A watery world.

In this photograph, you can see water as liquid (which covers three-quarters of Earth's surface), ice (in the form of snow), and vapor (as steam).

Structure/Function: Water

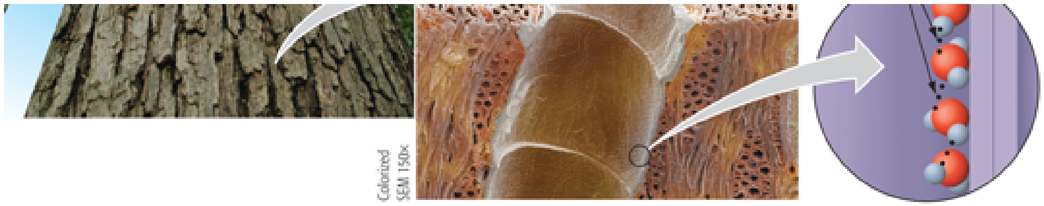
The unique properties of water on which all life on Earth depends are a prime example of one of biology's overarching themes: the relationship of structure and function. The structure of water molecules—the polarity and the hydrogen bonding that results (see Figure 2.8)—explains most of water's life-supporting functions. We'll explore four of those properties here: the cohesive nature of water, the ability of water to moderate temperature, the biological significance of ice floating, and the versatility of water as a solvent.

The Cohesion of Water

Water molecules stick together as a result of hydrogen bonding. In a drop of water, any particular set of hydrogen bonds lasts for only a few trillionths of a second, yet at any instant, huge numbers of hydrogen bonds exist between molecules of liquid water. 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 from their roots to their leaves (Figure 2.10).

Figure 2.10 Cohesion and water transport in plants.

The evaporation of water from leaves pulls water upward from the roots through microscopic tubes in the trunk of the tree. Because of cohesion, the pulling force is relayed through the tubes all the way down to the roots. As a result, water rises against the force 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 (Figure 2.11). Other liquids have much weaker surface tension; an insect, for example, could not walk on the surface of a cup of gasoline (which is why gardeners sometimes use gas to drown bugs removed from flower bushes).

Figure 2.11 A raft spider walking on water.

The cumulative strength of hydrogen bonds among water molecules allows this spider to walk on pond water without breaking the surface.

How Water Moderates Temperature

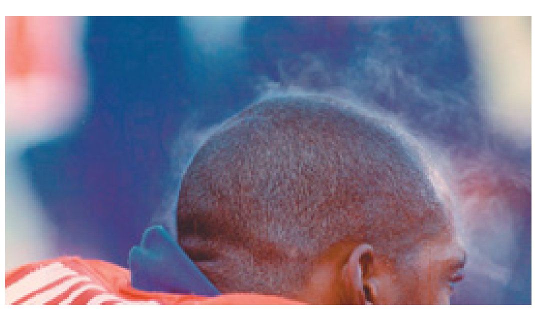
If you've ever burned your finger on a metal pot while waiting for the water in it to boil, you know that water heats up much more slowly than metal. In fact, because of hydrogen bonding, water has a stronger resistance to temperature change than most other substances.

When water is heated, the heat energy first disrupts hydrogen bonds and then makes water molecules jostle around faster. The temperature of the water doesn't go up until the water molecules start to speed up. Because heat is first used to break hydrogen bonds rather than raise the temperature, water absorbs and stores a large amount of heat while warming up only a few degrees. Conversely, when water cools, hydrogen bonds form, a process that releases heat. Thus, water can release a relatively large amount of heat to the surroundings while the water temperature drops only slightly.

Earth's giant water supply—the oceans, seas, lakes, and rivers—enables temperatures on the planet to stay within limits that permit life by storing a huge amount of heat from the sun during warm periods and giving off heat that warms the air during cold periods. 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. You may have noticed that the water temperature at the beach fluctuates much less than the air temperature.

Another way that water moderates temperature is by **evaporative cooling**. When a substance evaporates (changes from a liquid to a gas), the surface of the liquid that remains cools down. This occurs because the molecules with the greatest energy (the "hottest" ones) tend to vaporize first. Think of it like this: If the five fastest runners on your track team quit school, it would lower the average speed of the remaining team. Evaporative cooling helps prevent some land-dwelling creatures from overheating; it's why sweating helps you dissipate excess body heat (Figure 2.12). And the expression "It's not the heat, it's the humidity" has its basis in the difficulty of sweating when the air is already saturated with water vapor.

Figure 2.12 Sweating as a mechanism of evaporative cooling.

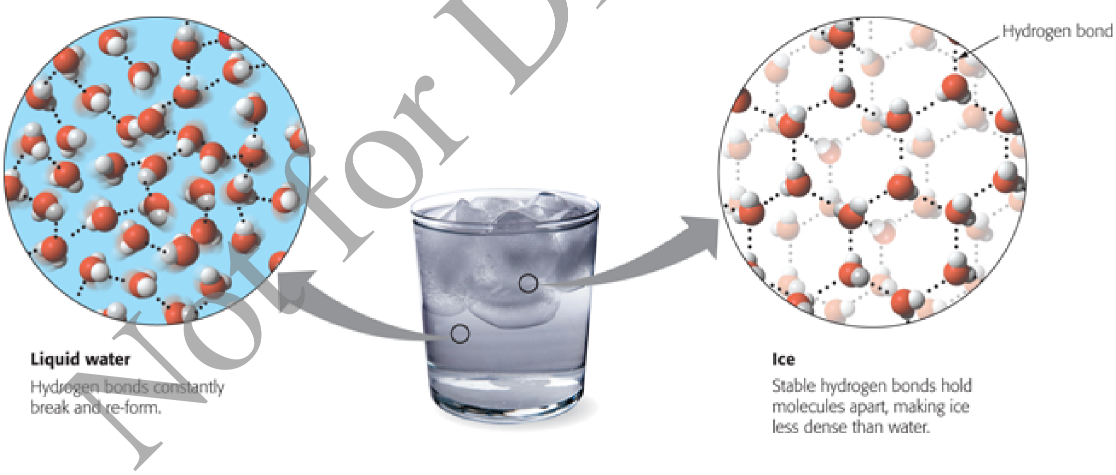


The Biological Significance of Ice Floating

When most liquids get cold, their molecules move closer together. If the temperature is cold enough, the liquid freezes and becomes a solid. Water, however, behaves differently. When water molecules get cold enough, they move apart, with each molecule staying at "arm's length" from its neighbors, forming ice. A chunk of ice floats because it is less dense than the liquid water in which it is floating. Floating ice is a consequence of hydrogen bonding. In contrast to the short-lived and constantly changing hydrogen bonds in liquid water, those in solid ice last longer, with each molecule bonded to four neighbors. As a result, ice is a spacious crystal (Figure 2.13).

Figure 2.13 Why ice floats.

Compare the tightly packed molecules in liquid water with the spaciously arranged molecules in the ice crystal. The less dense ice floats atop the denser water.



How does floating ice help support life on Earth? When a deep body of water cools and a layer of ice forms on top, the floating ice acts as an insulating "blanket" over the liquid water, allowing life to persist under the frozen surface. But imagine what would happen if ice were denser than water: Ice would sink during winter. All ponds, lakes, and even the oceans would eventually freeze solid without the insulating protection of the top layer of ice. Then, during summer, only the upper few inches of the oceans would thaw. It's hard to imagine life persisting under such conditions.

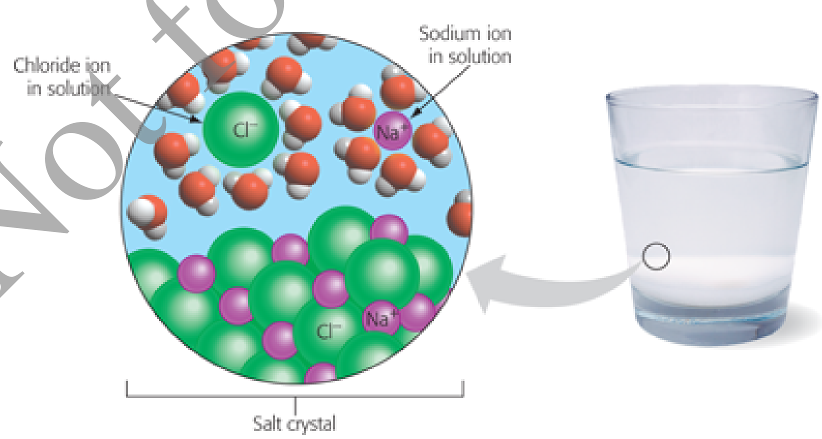
Water as the Solvent of Life

If you've ever stirred sugar into coffee or added salt to soup, you know that you can dissolve sugar or salt in water. This results in a mixture known as a **solution**, a liquid consisting of a homogeneous mixture of two or more substances. The dissolving agent is called the **solvent**, and any substance that is dissolved is called a **solute**. When water is the solvent, the resulting solution is called an **aqueous solution**. The fluids of organisms are aqueous solutions. For example, tree sap is an aqueous solution consisting of sugar and minerals dissolved in water.

Water can dissolve an enormous variety of solutes necessary for life, providing a medium for chemical reactions. For example, water can dissolve salt ions, as shown in Figure 2.14. Each ion becomes surrounded by oppositely charged regions of water molecules. Solutes that are polar molecules, such as sugars, dissolve by orienting locally charged regions of their molecules toward water molecules in a similar way.

Figure 2.14 A crystal of table salt (NaCl) dissolving in water.

As a result of electrical charge attractions, H_2O molecules surround the sodium and chloride ions, dissolving the crystal in the process.



We have discussed four special properties of water, each a consequence of water's unique chemical structure. Next, we'll look at aqueous solutions in more detail.

Checkpoint

1. Explain why, if you pour very carefully, you can actually "stack" water slightly above the rim of a cup.
2. Explain why ice floats.

   Answers: **1.** *Surface tension due to water's cohesion will keep the water from spilling over.* **2.** *Ice is less dense than liquid water because the more stable hydrogen bonds lock the molecules into a spacious crystal.*

Acids, Bases, and pH

In aqueous solutions, most of the water molecules are intact. However, some of the water molecules break apart into hydrogen ions (H^+) and hydroxide ions (OH^-). A balance of these two highly reactive ions is critical for the proper functioning of chemical processes within organisms.

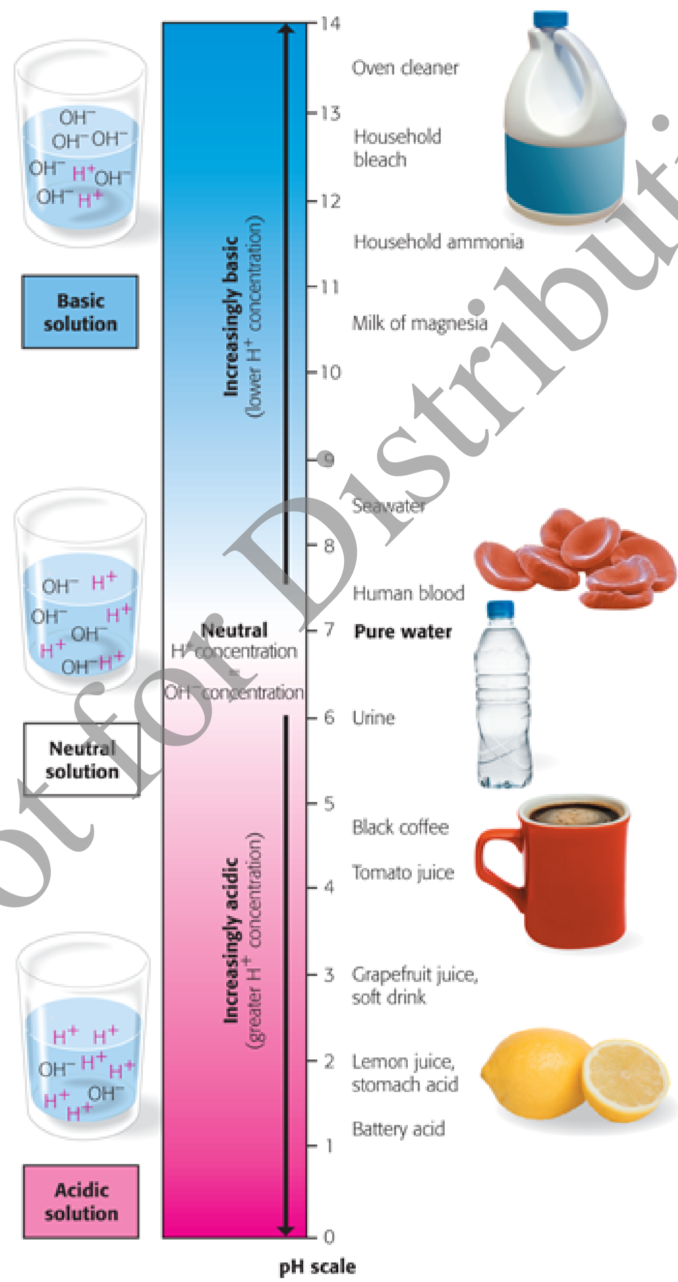
A chemical compound that releases H^+ to a solution is called an **acid**. One example of a strong acid is hydrochloric acid (HCl), the acid in your stomach that aids in digestion of food. In solution, HCl breaks apart into the ions H^+ and Cl^-. A **base** (or alkali) is a compound that accepts H^+ and removes them from solution. Some bases, such as sodium hydroxide (NaOH), do this by releasing OH^-, which combines with H^+ to form H_2O.

Lemon juice has roughly the same acidity as the food-digesting chemicals in your stomach.

To describe the acidity of a solution,chemists use the **pH scale**, a measure of the hydrogen ion (H^+) concentration in a solution. The scale ranges from 0 (most acidic) to 14 (most basic). Each pH unit represents a tenfold change in the concentration of H^+ (Figure 2.15). For example, lemon juice at pH 2 has 100 times more H^+ than an equal amount of tomato juice at pH 4. Aqueous solutions that are neither acidic nor basic (such as pure water) are said to be neutral; they have a pH of 7. They do contain some H^+ and OH^-, but the concentrations of the two ions are equal. The pH of the solution inside most living cells is close to 7.

Figure 2.15 **The pH scale.**

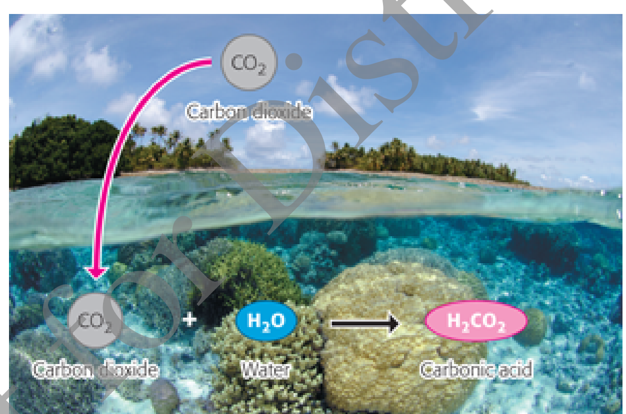
A solution having a pH of 7 is neutral, meaning that its H^+ and OH^- concentrations are equal. The lower the pH below 7, the more acidic the solution, or the greater its excess of H^+ compared with OH^-. The higher the pH above 7, the more basic the solution, or the greater the deficiency of H^+ relative to OH^-.



Even a slight change in pH can be harmful to an organism because the molecules in cells are extremely sensitive to H^+ and OH^- concentrations. Biological fluids contain **buffers**, substances that minimize changes in pH by accepting H^+ when that ion is in excess and donating H^+ when it is depleted. For example, buffer in contact lens solution helps protect the surface of the eye from potentially painful changes in pH. This buffering process, however, is not foolproof, and changes in environmental pH can profoundly affect ecosystems. For example, about 25% of the carbon dioxide (CO_2) generated by people (primarily by burning fossil fuels) is absorbed by the oceans. When CO_2 dissolves in seawater, it reacts with water to form carbonic acid (Figure 2.16), which lowers ocean pH. The resulting ocean acidification can greatly change marine environments. Oceanographers have calculated that the pH of the ocean is lower now than at any time in the past 420,000 years, and it is continuing to drop.

Figure 2.16 Ocean acidification by atmospheric CO_2.

After dissolving in seawater, CO_2 reacts to form carbonic acid. This acid then undergoes further chemical reactions that disrupt coral growth. Such acidification can cause drastic changes in important marine ecosystems.



The effects of ocean acidification—including coral bleaching and changes in metabolism among a wide variety of sea creatures—are daunting reminders that the chemistry of life is linked to the chemistry of the environment. It reminds us, too, that chemistry happens on a global scale, because industrial processes in one region of the world often cause ecosystem changes in another part of the world.

Checkpoint

Compared with a solution of pH 8, the same volume of a solution at pH 5 has

Compared with a solution of pH 8, the same volume of a solution at pH 5 has ______ times more hydrogen ions (H^+). This second solution is considered a(n) ______.

Answer: *1,000; acid*

Radioactivity: Evolution Connection

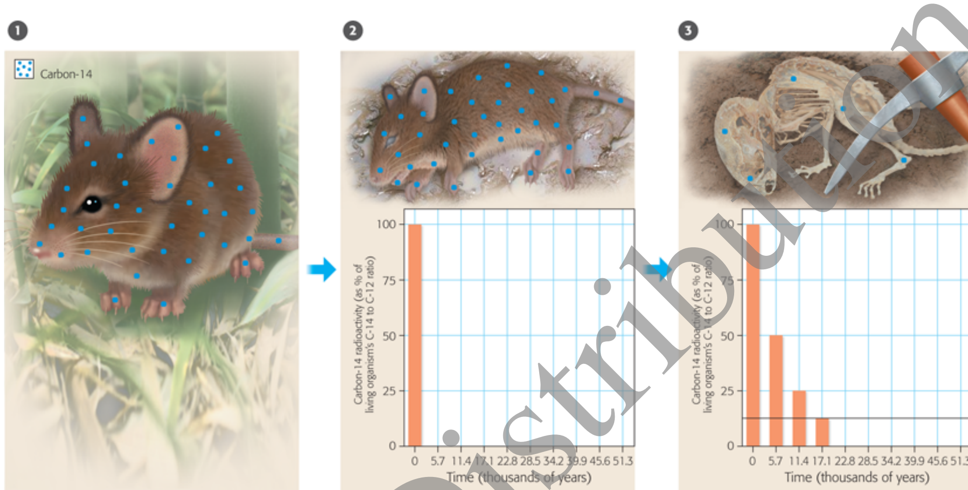
Radioactivity as an Evolutionary Clock

Throughout this chapter, we have highlighted ways—both helpful and harmful—that radioactivity can affect the health of living organisms. In addition to detecting and treating diseases, another helpful application of radioactivity involves the natural process of radioactive decay, which can be used to obtain important data about the evolutionary history of life on Earth.

Fossils—both preserved imprints and remains of dead organisms—are reliable chronological records of life because we can determine their ages through radiometric dating (Figure 2.17), which is based on the decay of radioactive isotopes. For example, carbon-14 is a radioactive isotope with a half-life of 5,700 years. It is present in trace amounts in the environment. 1 A living organism assimilates the different isotopes of an element in proportions that reflect their relative abundances in the environment. In this example, carbon-14 is taken up in trace quantities, along with much larger quantities of the more common carbon-12. 2 At the time of death, the organism ceases to take in carbon from the environment. From this moment forward, the amount of carbon-14 relative to carbon-12 in the fossil declines: The carbon-14 in the body decays into carbon-12, but no new carbon-14 is added. Because the half-life of carbon-12 is known, the ratio of the two isotopes (carbon-14 to carbon-12) is a reliable indicator of the age of the fossil. In this case, it takes 5,700 years for half of the radioactive carbon-14 to decay; half of the remainder is present after another 5,700 years; and so on. 3 A fossil's age can be estimated by measuring the ratio of the two isotopes to learn how many half-life reductions have occurred since it died. For example, if the ratio of carbon-14 to carbon-12 in this fossil was found to be ⅛th that of the environment, this fossil would be about 17,100 (5,700 × 3 half-lives) years old.

Figure 2.17 **Radiometric dating.**

Living organisms incorporate the carbon-14 isotope (represented here with blue dots). But no new carbon-14 is introduced once an organism dies, and the carbon-14 that remains slowly decays into carbon-12. By measuring the amount of carbon-14 in a fossil, scientists can estimate its age.



Using such techniques, scientists can estimate the ages of fossils from around the world and place them in an ordered sequence called the fossil record. The fossil record is one of the most important and convincing sets of evidence that led Charles Darwin to formulate the theory of natural selection (as you'll see later, in Chapter 14). Every time a new fossil is dated, it can be placed within the vast history of life on Earth.