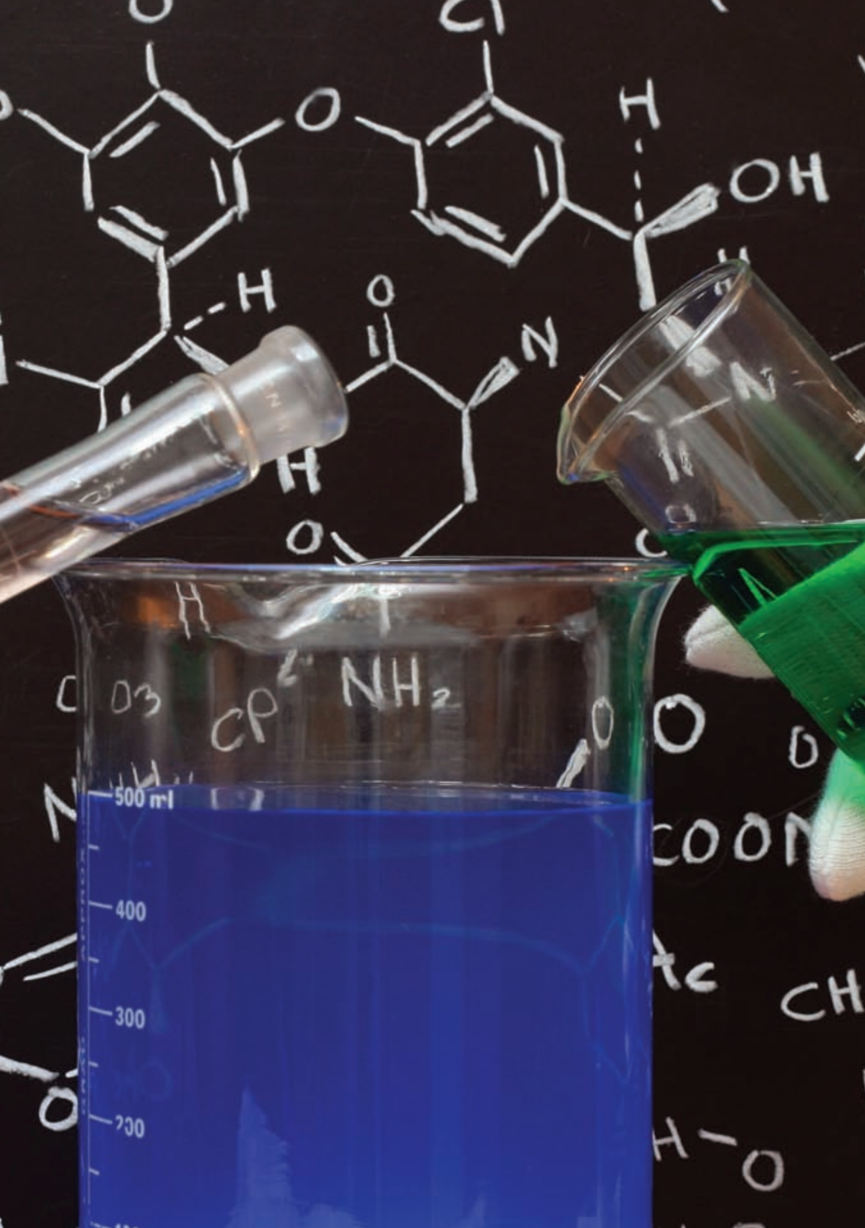


CHAPTER 2

The Nature of Molecules and the Properties of Water

Chapter Contents

Introduction

About 12.5 billion years ago (BYA), an enormous explosion probably signaled the beginning of the universe. This explosion started a process of star building and planetary formation that eventually led to the formation of Earth, about 4.5 BYA. Around 3.5 BYA, life began on Earth and started to diversify. To understand the nature of life on Earth, we first need to understand the nature of the matter that forms the building blocks of all life.

The earliest speculations about the world around us included this most basic question, "What is it made of?" The ancient Greeks recognized that larger things may be built of smaller parts. This concept was formed into a solid experimental scientific idea in the early 20th century, when physicists began trying to break atoms apart. From those humble beginnings to the huge particle accelerators used by the modern physicists of today, the picture of the atomic world emerges as fundamentally different from the tangible, macroscopic world around us.

To understand how living systems are assembled, we must first understand a little about atomic structure, about how atoms can be linked together by chemical bonds to make molecules, and about the ways in which these small molecules are joined together to make larger molecules, until finally we arrive at the structures of cells and then of organisms. Our study of life on Earth therefore begins with physics and chemistry. For many of you, this chapter will be a review of material encountered in other courses.

2.1 The Nature of Atoms

Learning Outcomes

1. *Define an element based on its composition.*
2. *Describe the relationship between atomic structure and chemical properties.*
3. *Explain where electrons are found in an atom.*

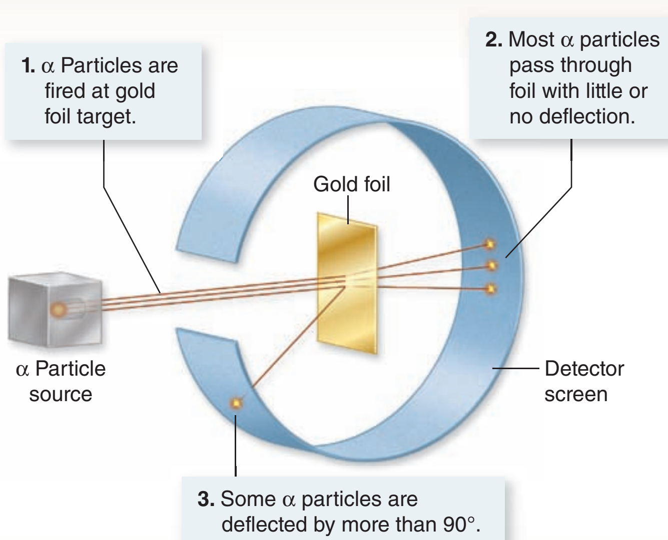
Any substance in the universe that has mass and occupies space is defined as *matter.* All matter is composed of extremely small particles called **atoms.** Because of their size, atoms are difficult to study. Not until early in the 20th century did scientists carry out the first experiments revealing the physical nature of atoms (figure 2.1).

SCIENTIFIC THINKING

Hypothesis: *Atoms are composed of diffuse positive charge with embedded negative charge (electrons).*

Prediction: *If alpha (α) particles, which are helium nuclei, are shot at a thin foil of gold, the α particles will not be deflected much by the diffuse positive charge or by the light electrons.*

Test: *α Particles are shot at a thin sheet of gold foil surrounded by a detector screen, which shows flashes of light when hit by the particles.*



Result: *Most particles are not deflected at all, but a small percentage of particles are deflected at angles of 90° or more.*

Conclusion: *The hypothesis is not supported. The large deflections observed led to a view of the atom as composed of a very small central region containing positive charge (the nucleus) surrounded by electrons.*

Further Experiments: *How does the Bohr atom with its quantized energy for electrons extend this model?*

Figure 2.1 Rutherford scattering experment. Large-angle scattering of α particles led Rutherford to propose the existence of the nucleus.

Atomic structure includes a central nucleus and orbiting electrons

Objects as small as atoms can be "seen" only indirectly, by using complex technology such as tunneling microscopy (figure 2.2). We now know a great deal about the complexities of atomic structure, but the simple view put forth in 1913 by the Danish physicist Niels Bohr provides a good starting point for understanding atomic theory. Bohr proposed that every atom possesses an orbiting cloud of tiny subatomic particles called *electrons* whizzing around a core, like the planets of a miniature solar system. At the center of each atom is a small, very dense nucleus formed of two other kinds of subatomic particles: *protons* and *neutrons* (figure 2.3).

Atomic number

Different atoms are defined by the number of protons, a quantity called the *atomic number*. Atoms with the same atomic number (that is, the same number of protons) have the same chemical properties and are said to belong to the same element. Formally speaking, an *element* is any substance that cannot be broken down to any other substance by ordinary chemical means.

Within the nucleus, the cluster of protons and neutrons is held together by a force that works only over short, subatomic distances. Each proton carries a positive (+) charge, and each neutron has no charge. Each electron carries a negative (–) charge. Typically, an atom has one electron for each proton and is thus electrically neutral. The chemical behavior of an atom is due to the number and configuration of electrons, as we will see later in this section.

Atomic mass

The terms *mass* and *weight* are often used interchangeably, but they have slightly different meanings. *Mass* refers to the amount of a substance, but *weight* refers to the force gravity exerts on a substance. An object has the same mass whether it is on the Earth or the Moon, but its weight will be greater on the Earth because the Earth's gravitational force is greater than the Moon's. The *atomic mass* of an atom is equal to the sum of the masses of its protons and neutrons. Atoms that occur naturally on Earth contain from 1 to 92 protons and up to 146 neutrons.

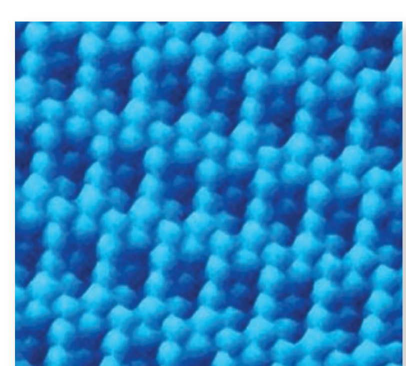


Figure 2.2 Scanning-tunneling microscope image. The scanning-tunneling microscope is a nonoptical way of imaging that allows atoms to be visualized. This image shows a lattice of oxygen atoms *(dark blue)* on a rhodium crystal *(light blue).*

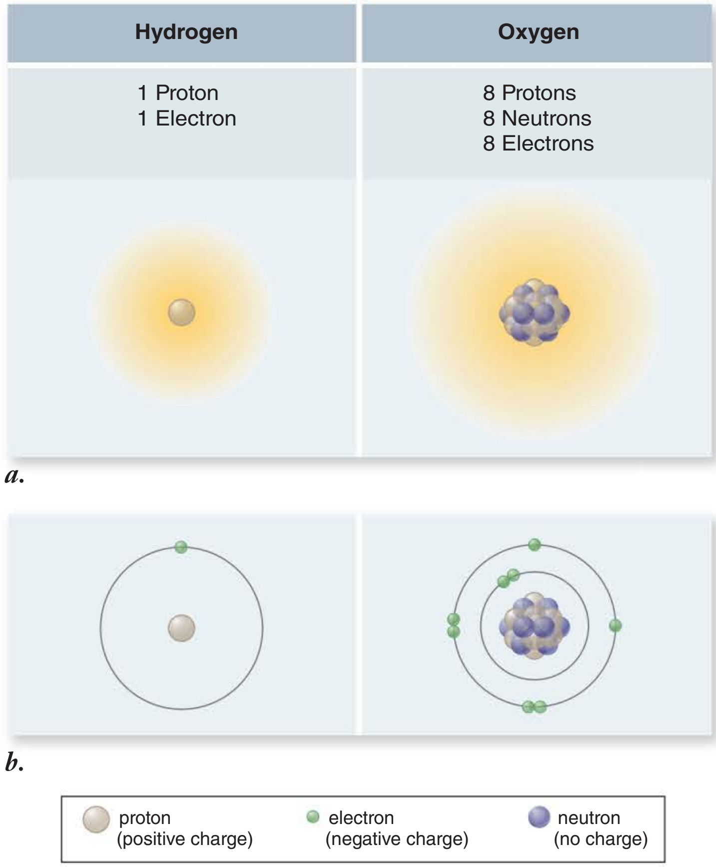


Figure 2.3 Basic structure of atoms. All atoms have a nucleus consisting of protons and neutrons, except hydrogen, the smallest atom, which usually has only one proton and no neutrons in its nucleus. Oxygen typically has eight protons and eight neutrons in its nucleus. In the simple "Bohr model" of atoms pictured here, electrons spin around the nucleus at a relatively far distance. ***a.*** Atoms are depicted as a nucleus with a cloud of electrons (not shown to scale). ***b.*** The electrons are shown in discrete energy levels. These are described in greater detail in the text.

The mass of atoms and subatomic particles is measured in units called *daltons.* To give you an idea of just how small these units are, note that it takes 602 million million billion (6.02×10^{23}) daltons to make 1 gram (g). A proton weighs approximately 1 dalton (actually 1.007 daltons), as does a neutron (1.009 daltons). In contrast, electrons weigh only 1/1840 of a dalton, so they contribute almost nothing to the overall mass of an atom.

Electrons

The positive charges in the nucleus of an atom are neutralized, or counterbalanced, by negatively charged electrons, which are located in regions called **orbitals** that lie at varying distances around the nucleus. Atoms with the same number of protons and electrons are electrically neutral—that is, they have no net charge, and are therefore called *neutral atoms.*

Electrons are maintained in their orbitals by their attraction to the positively charged nucleus. Sometimes other forces overcome this attraction, and an atom loses one or more electrons. In other cases, atoms gain additional electrons. Atoms in which the number of electrons does not equal the number of protons are known as *ions,* and they are charged particles. An atom having more protons than electrons has a net positive charge and is called a **cation.** For example, an atom of sodium (Na) that has lost one electron becomes a sodium ion (Na^+), with a charge of +1. An atom having fewer protons than electrons carries a net negative charge and is called an **anion.** A chlorine atom (Cl) that has gained one electron becomes a chloride ion (Cl^-), with a charge of −1.

Isotopes

Although all atoms of an element have the same number of protons, they may not all have the same number of neutrons. Atoms of a single element that possess different numbers of neutrons are called **isotopes** of that element.

Most elements in nature exist as mixtures of different isotopes. Carbon (C), for example, has three isotopes, all containing six protons (figure 2.4). Over 99% of the carbon found in nature exists as an isotope that also contains six neutrons. Because the total mass of this isotope is 12 daltons (6 from protons plus 6 from neutrons), it is referred to as carbon-12 and is symbolized ^{12}C. Most of the rest of the naturally occurring carbon is carbon-13, an isotope with seven neutrons. The rarest carbon isotope is carbon-14, with eight neutrons. Unlike the other two isotopes, carbon-14 is unstable: This means that its nucleus tends to break up into elements with lower atomic numbers. This nuclear breakup, which emits a significant amount of energy, is called *radioactive decay,* and isotopes that decay in this fashion are **radioactive isotopes.**

Some radioactive isotopes are more unstable than others, and therefore they decay more readily. For any given isotope, however, the rate of decay is constant. The decay time is usually

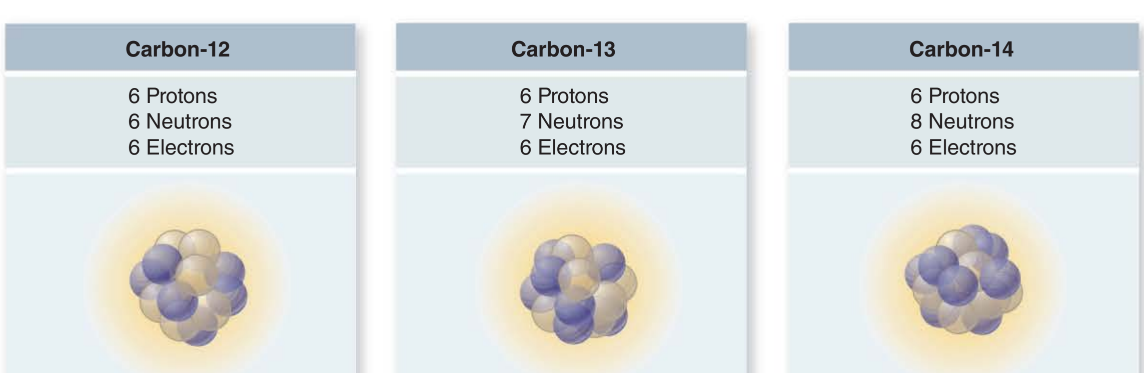


Figure 2.4 The three most abundant isotopes of carbon. Isotopes of a particular element have different numbers of neutrons.

expressed as the *half-life,* the time it takes for one-half of the atoms in a sample to decay. Carbon-14, for example, often used in the carbon dating of fossils and other materials, has a half-life of 5730 years. A sample of carbon containing 1 g of carbon-14 today would contain 0.5 g of carbon-14 after 5730 years, 0.25 g 11,460 years from now, 0.125 g 17,190 years from now, and so on. By determining the ratios of the different isotopes of carbon and other elements in biological samples and in rocks, scientists are able to accurately determine when these materials formed.

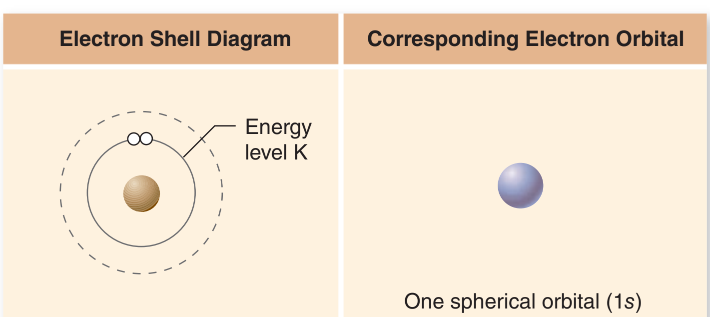
Radioactivity has many useful applications in modern biology. Radioactive isotopes are one way to label, or "tag," a specific molecule and then follow its progress, either in a chemical reaction or in living cells and tissue. The downside, however, is that the energetic subatomic particles emitted by radioactive substances have the potential to severely damage living cells, producing genetic mutations and, at high doses, cell death. Consequently, exposure to radiation is carefully controlled and regulated. Scientists who work with radioactivity follow strict handling protocols and wear radiation-sensitive badges to monitor their exposure over time to help ensure a safe level of exposure.

Electrons determine the chemical behavior of atoms

The key to the chemical behavior of an atom lies in the number and arrangement of its electrons in their orbitals. The Bohr model of the atom shows individual electrons as following distinct circular orbits around a central nucleus. The trouble with this simple picture is that it doesn't reflect reality. Modern physics indicates that we cannot pinpoint the position of any individual electron at any given time. In fact, an electron could be anywhere, from close to the nucleus to infinitely far away from it.

A particular electron, however, is more likely to be in some areas than in others. An orbital is defined as the area around a nucleus where an electron is most likely to be found. These orbitals represent probability distributions for electrons—that is, regions more likely to contain an electron. Some electron orbitals near the nucleus are spherical (*s* orbitals), whereas others are dumbbell-shaped (*p* orbitals) (figure 2.5). Still other orbitals, farther away from the nucleus, may have different shapes. Regardless of its shape, no orbital can contain more than two electrons.

Almost all of the volume of an atom is empty space. This is because the electrons are usually far away from the nucleus, relative to its size. If the nucleus of an atom were the size of a golf ball, the orbit of the nearest electron would be a mile away. Consequently, the nuclei of two atoms never come close enough in nature to interact with each other. It is for this reason that an atom's electrons, not its protons or neutrons, determine its chemical behavior, and it also explains why the isotopes of an element, all of which have the same arrangement of electrons, behave the same way chemically.



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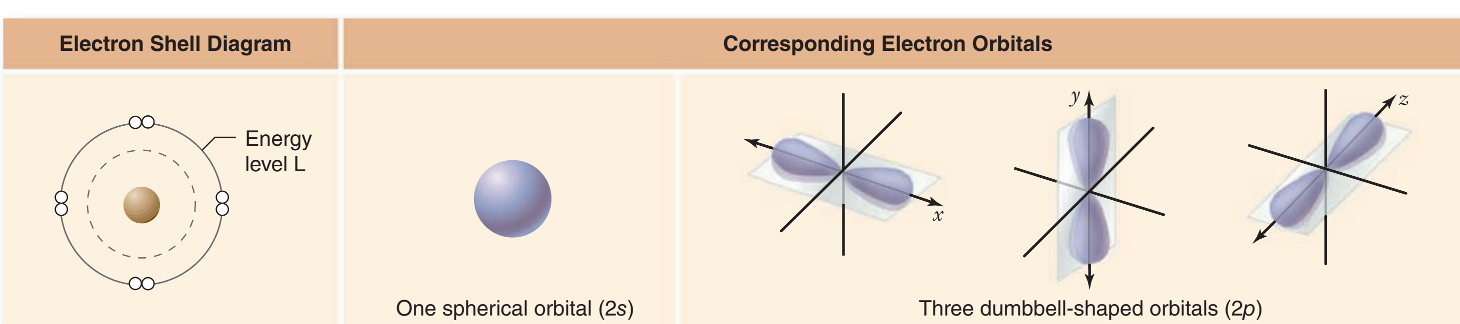

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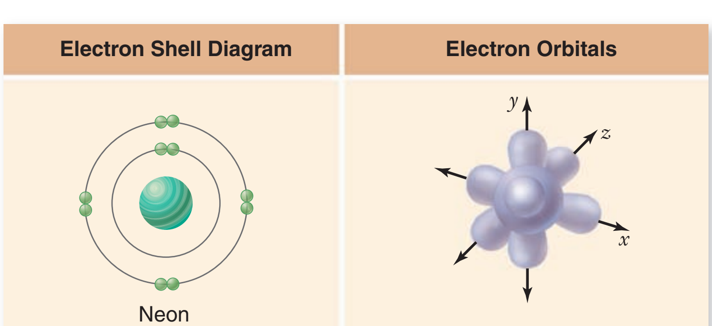


c.

Figure 2.5 Electron orbitals. ***a.*** The lowest energy level, or electron shell—the one nearest the nucleus—is level K. It is occupied by a single *s* orbital, referred to as 1*s*. ***b.*** The next highest energy level, L, is occupied by four orbitals: one *s* orbital (referred to as the 2*s* orbital) and three *p* orbitals (each referred to as a 2*p* orbital). Each orbital holds two paired electrons with opposite spin. Thus, the K level is populated by two electrons, and the L level is populated by a total of eight electrons. ***c.*** The neon atom shown has the L and K energy levels completely filled with electrons and is thus unreactive.

Atoms contain discrete energy levels

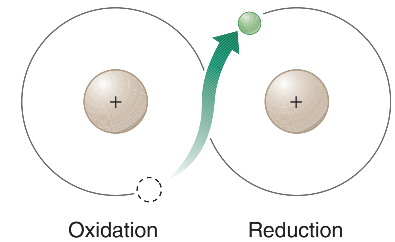
Because electrons are attracted to the positively charged nucleus, it takes work to keep them in their orbitals, just as it takes work to hold a grapefruit in your hand against the pull of gravity. The formal definition of energy is the ability to do work.

The grapefruit held above the ground is said to possess *potential energy* because of its position. If you release it, the grapefruit falls, and its potential energy is reduced. On the other hand, if you carried the grapefruit to the top of a building, you would increase its potential energy. Electrons also have a potential energy that is related to their position. To oppose the attraction of the nucleus and move the electron to a more distant orbital requires an input of energy, which results in an electron with greater potential energy. The chlorophyll that makes plants green captures energy from light during photosynthesis in this way. As you'll see in chapter 8—light energy excites electrons in the chlorophyll molecule. Moving an electron closer to the nucleus has the opposite effect: Energy is released, usually as radiant energy (heat or light), and the electron ends up with less potential energy (figure 2.6).

One of the initially surprising aspects of atomic structure is that electrons within the atom have discrete **energy levels.** These discrete levels correspond to quanta (singular, quantum), which means specific amount of energy. To use the grapefruit analogy again, it is as though a grapefruit could only be raised to particular floors of a building. Every atom exhibits a ladder of potential energy values, a discrete set of orbitals at particular energetic "distances" from the nucleus.

Because the amount of energy an electron possesses is related to its distance from the nucleus, electrons that are the same distance from the nucleus have the same energy, even if they occupy different orbitals. Such electrons are said to occupy the same energy level. The energy levels are denoted with letters K, L, M, and so on (figure 2.6). Be careful not to confuse energy levels, which are drawn as rings to indicate an electron's *energy,* with orbitals, which have a variety of three-dimensional shapes and indicate an electron's most likely *location.* Electron orbitals are arranged so that as they are filled, this fills each energy level in successive order. This filling of orbitals and energy levels is what is responsible for the chemical reactivity of elements.

During some chemical reactions, electrons are transferred from one atom to another. In such reactions, the loss of an electron is called **oxidation,** and the gain of an electron is called *reduction.*



Notice that when an electron is transferred in this way, it keeps its energy of position. In organisms, chemical energy is stored in high-energy electrons that are transferred from one atom to another in reactions involving oxidation and reduction (described in chapter 7). When the processes of oxidation and reduction are coupled, which often happens, one atom or molecule is oxidized, while another is reduced in the same reaction. We call these combinations *redox reactions.*

Learning Outcomes Review 2.1

An atom consists of a nucleus of protons and neutrons surrounded by a cloud of electrons. For each atom, the number of protons is the atomic number; atoms with the same atomic number constitute an element. Atoms of a single element that have different numbers of neutrons are called isotopes. Electrons, which determine the chemical behavior of an element, are located about a nucleus in orbitals representing discrete energy levels. No orbital can contain more than two electrons, but each energy level consists of multiple orbitals, and thus contains many electrons with the same energy.

- ***If the number of protons exceeds the number of neutrons, is the charge on the atom positive or negative?***
- ***If the number of protons exceeds electrons?***

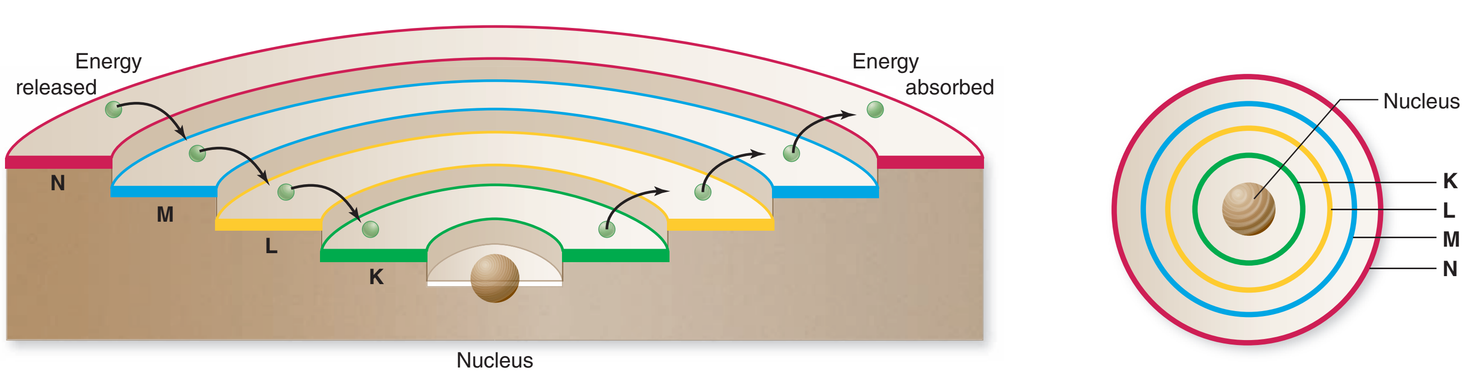


Figure 2.6 Atomic energy levels. Electrons have energy of position. When an atom absorbs energy, an electron moves to a higher energy level, farther from the nucleus. When an electron falls to lower energy levels, closer to the nucleus, energy is released. The first two energy levels are the same as shown in figure 2.5.

2.2 Elements Found in Living Systems

Learning Outcomes

1. *Relate atomic structure to the periodic table of the elements.*
2. *List the important elements found in living systems.*

Ninety elements occur naturally, each with a different number of protons and a different arrangement of electrons. When the 19th-century Russian chemist Dmitri Mendeleev arranged the known elements in a table according to their atomic number, he discovered one of the great generalizations of science: The elements exhibit a pattern of chemical properties that repeats itself in groups of eight. This periodically repeating pattern lent the table its name: the periodic table of elements (figure 2.7).

The periodic table displays elements according to atomic number and properties

The eight-element periodicity that Mendeleev found is based on the interactions of the electrons in the outermost energy level of the different elements. These electrons are called **valence electrons,** and their interactions are the basis for the elements' differing chemical properties. For most of the atoms important to life, the outermost energy level can contain no more than eight electrons; the chemical behavior of an element reflects how many of the eight positions are filled. Elements possessing all eight electrons in their outer energy level (two for helium) are *inert,* or nonreactive. These elements, which include helium (He), neon (Ne), argon (Ar), and so on, are called the *noble gases*. In sharp contrast, elements with seven electrons (one fewer than the maximum number of eight) in their outer energy level, such as fluorine (F), chlorine (Cl), and bromine (Br), are highly reactive. They tend to gain the extra electron needed to fill the energy level. Elements with only one electron in their outer energy level, such as lithium (Li), sodium (Na), and potassium (K), are also very reactive, but they tend to lose the single electron in their outer level.

Mendeleev's periodic table leads to a useful generalization, the **octet rule,** or *rule of eight* (Latin *octo*, "eight"): Atoms tend to establish completely full outer energy levels. For the main group elements of the periodic table, the rule of eight is accomplished by one filled *s* orbital and three filled *p* orbitals (figure 2.8). The exception to this is He, in the first row, which needs only two electrons to fill the 1*s* orbital. Most chemical behavior of biological interest can be predicted quite accurately from this simple rule, combined with the tendency of atoms to balance positive and negative charges. For instance, you read earlier that sodium ion (Na^+) has lost an electron, and chloride ion (Cl^-) has gained an electron. In the section 2.3, we describe how these ions react to form table salt.

Of the 90 naturally occurring elements on Earth, only 12 (C, H, O, N, P, S, Na, K, Ca, Mg, Fe, Cl) are found in living systems in more than trace amounts (0.01% or higher). These elements all have atomic numbers less than 21, and thus, have low atomic masses. Of these 12, the first 4 elements (carbon, hydrogen, oxygen, and nitrogen) constitute 96.3% of the weight of your body. The majority of molecules that make up your body (other than water) are compounds of carbon, which we call *organic* compounds.

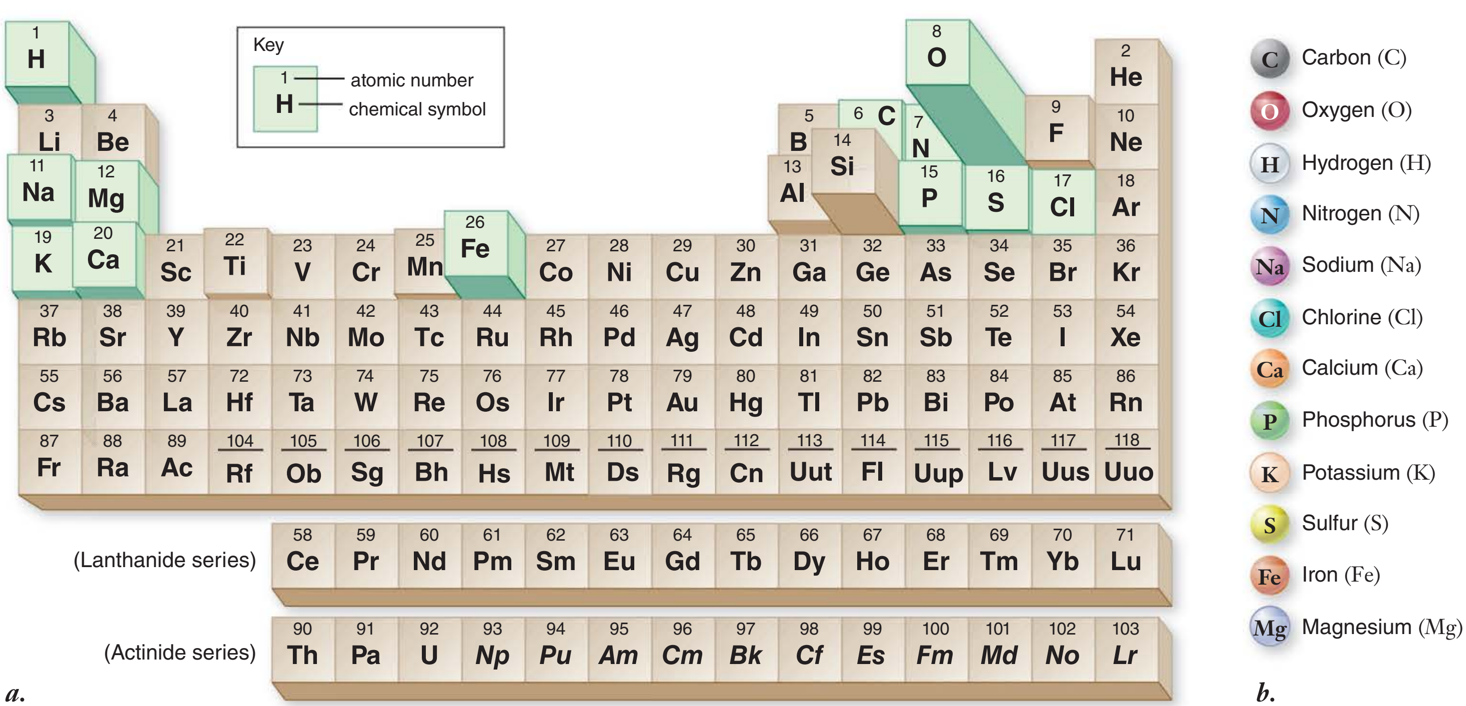


Figure 2.7 Periodic table of the elements. ***a.*** In this representation, the frequency of elements that occur in the Earth's crust is indicated by the height of the block. Elements shaded in green are found in living systems in more than trace amounts. ***b.*** Common elements found in living systems are shown in colors that will be used throughout the text.

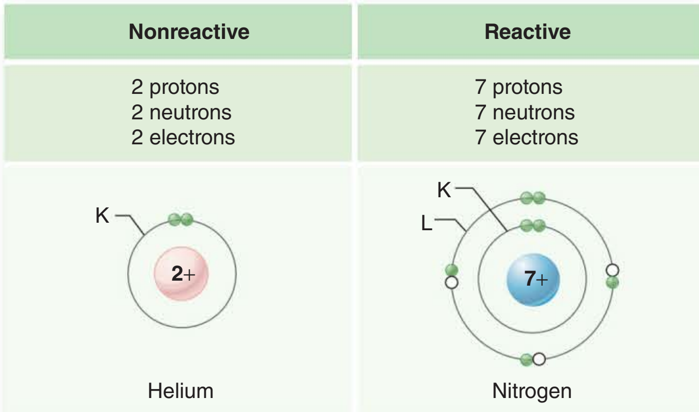


Figure 2.8 Electron energy levels for helium and nitrogen. Green balls represent electrons, blue ball represents the nucleus with number of protons indicated by number of (+) charges. Note that the helium atom has a filled K shell and is thus unreactive, whereas the nitrogen atom has five electrons in the L shell, three of which are unpaired, making it reactive.

These organic compounds contain primarily these four elements (CHON), explaining their prevalence in living systems. Some trace elements, such as zinc (Zn) and iodine (I), play crucial roles in living processes even though they are present in tiny amounts. Iodine deficiency, for example, can lead to enlargement of the thyroid gland, causing a bulge at the neck called a goiter.

Learning Outcomes Review 2.2

The periodic table shows the elements in terms of atomic number and repeating chemical properties. Only 12 elements are found in significant amounts in living organisms: C, H, O, N, P, S, Na, K, Ca, Mg, Fe, and Cl.

- *Why are the noble gases more stable than other elements in the periodic table?*

2.3 The Nature of Chemical Bonds

Learning Outcomes

1. *Predict which elements are likely to form ions.*
2. *Explain how molecules are formed from atoms joined by covalent bonds.*
3. *Contrast polar and nonpolar covalent bonds.*

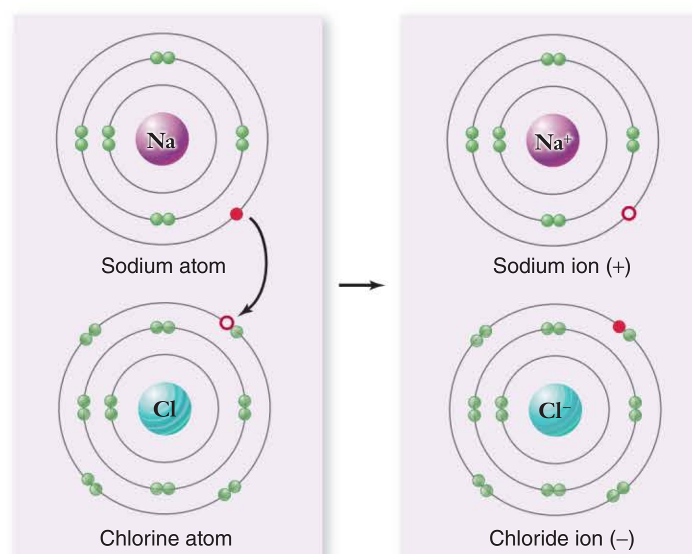
A group of atoms held together by energy in a stable association is called a *molecule.* When a molecule contains atoms of more than one element, it is called a *compound.* The atoms in a molecule are joined by *chemical bonds;* these bonds can result when atoms with opposite charges attract each other (ionic bonds), when two atoms share one or more pairs of electrons (covalent bonds), or when atoms interact in other ways (table 2.1). We will start by examining *ionic bonds,* which form when atoms with opposite electrical charges (ions) attract.

TABLE 2.1 Bonds and Interactions

Name	Basis of Interaction	Strength
Covalent bond	Sharing of electron pairs	Strong
Ionic bond	Attraction of opposite charges	↑
Hydrogen bond	Sharing of H atom	
Hydrophobic interaction	Forcing of hydrophobic portions of molecules together in presence of polar substances	↓
van der Waals attraction	Weak attractions between atoms due to oppositely polarized electron clouds	Weak

Ionic bonds form crystals

Common table salt, the molecule sodium chloride (NaCl), is a lattice of ions in which the atoms are held together by ionic bonds (figure 2.9). Sodium has 11 electrons: 2 in the inner energy



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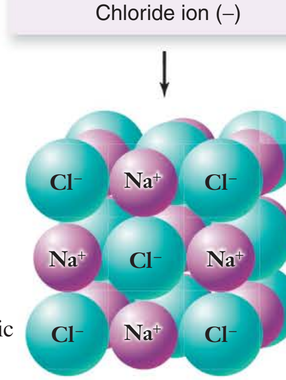


b. NaCl crystal

Figure 2.9 The formation of ionic bonds by sodium chloride. ***a.*** When a sodium atom donates an electron to a chlorine atom, the sodium atom is oxidized and the chlorine atom reduced. This produces a positively charged sodium ion, and a negatively charged chloride ion. ***b.*** The electrostatic attraction of oppositely charged ions leads to the formation of a lattice of Na^+ and Cl^-.

level (K), 8 in the next level (L), and 1 in the outer (valence) level (M). The single, unpaired valence electron has a strong tendency to join with another unpaired electron in another atom. A stable configuration can be achieved if the valence electron is lost to another atom that also has an unpaired electron. The loss of this electron results in the formation of a positively charged sodium ion, Na^+.

The chlorine atom has 17 electrons: 2 in the K level, 8 in the L level, and 7 in the M level. As you can see in the figure, one of the orbitals in the outer energy level has an unpaired electron (red circle). The addition of another electron fills that level and causes a negatively charged chloride ion, Cl^-, to form.

When placed together, metallic sodium and gaseous chlorine react swiftly and explosively, as the sodium atoms are oxidized, donating electrons to chlorine atoms, reducing them, and forming Na^+ and Cl^- ions. Because opposite charges attract, the Na^+ and Cl^- remain associated in an *ionic compound,* NaCl, which is electrically neutral. The electrical attractive force holding NaCl together, however, is not directed specifically between individual Na^+ and Cl^- ions, and no individual sodium chloride molecules form. Instead, the force exists between any one ion and *all* neighboring ions of the opposite charge. The ions aggregate in a crystal matrix with a precise geometry. Such aggregations are what we know as salt crystals. If a salt such as NaCl is placed in water, the electrical attraction of the water molecules disrupts the forces holding the ions in their crystal matrix, causing the salt to dissolve into a roughly equal mixture of free Na^+ and Cl^- ions.

Because living systems always include water, ions are more important than ionic crystals. Important ions in biological systems include Ca^{2+}, which is involved in cell signaling, K^+ and Na^+, which are involved in the conduction of nerve impulses.

Covalent bonds build stable molecules

Covalent bonds form when two atoms share one or more pairs of valence electrons. Consider gaseous hydrogen (H_2) as an example. Each hydrogen atom has an unpaired electron and an unfilled outer energy level; for these reasons, the hydrogen atom is unstable. However, when two hydrogen atoms are in close association, each atom's electron is attracted to both nuclei. In effect, the nuclei are able to share their electrons. The result is a diatomic (two-atom) molecule of hydrogen gas.

The molecule formed by the two hydrogen atoms is stable for three reasons:

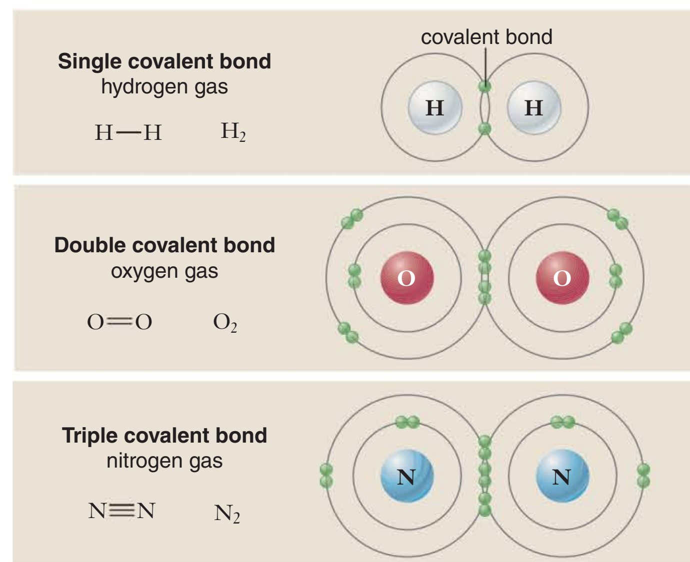
1. **It has no net charge.** The diatomic molecule formed as a result of this sharing of electrons is not charged because it still contains two protons and two electrons.
2. **The octet rule is satisfied.** Each of the two hydrogen atoms can be considered to have two orbiting electrons in its outer energy level. This state satisfies the octet rule, because each shared electron orbits both nuclei and is included in the outer energy level of both atoms.
3. **It has no unpaired electrons.** The bond between the two atoms also pairs the two free electrons.

Unlike ionic bonds, covalent bonds are formed between two individual atoms, giving rise to true, discrete molecules.

The strength of covalent bonds

The strength of a covalent bond depends on the number of shared electrons. Thus *double bonds,* which satisfy the octet rule by allowing two atoms to share two pairs of electrons, are stronger than *single bonds,* in which only one electron pair is shared. In practical terms, more energy is required to break a double bond than a single bond. The strongest covalent bonds are *triple bonds,* such as those that link the two nitrogen atoms of nitrogen gas molecules (N_2).

Covalent bonds are represented in chemical formulas as lines connecting atomic symbols. Each line between two bonded atoms represents the sharing of one pair of electrons. The *structural formulas* of hydrogen gas and oxygen gas are H—H and O═O, respectively, and their *molecular formulas* are H_2 and O_2. The structural formula for N_2 is N≡N.



Molecules with several covalent bonds

A vast number of biological compounds are composed of more than two atoms. An atom that requires two, three, or four additional electrons to fill its outer energy level completely may acquire them by sharing its electrons with two or more other atoms.

For example, the carbon atom (C) contains six electrons, four of which are in its outer energy level and are unpaired. To satisfy the octet rule, a carbon atom must form four covalent bonds. Because four covalent bonds may form in many ways, carbon atoms are found in many different kinds of molecules. CO_2 (carbon dioxide), CH_4 (methane), and C_2H_5OH (ethanol) are just a few examples.

Polar and nonpolar covalent bonds

Atoms differ in their affinity for electrons, a property called **electronegativity.** In general, electronegativity increases left to right across a row of the periodic table and decreases down the column. Thus the elements in the upper-right corner have the highest electronegativity.

For bonds between identical atoms, for example, between two hydrogen or two oxygen atoms, the affinity for electrons is obviously the same, and the electrons are equally shared. Such bonds are termed **nonpolar.** The resulting compounds (H_2 or O_2) are also referred to as nonpolar.

For atoms that differ greatly in electronegativity, electrons are not shared equally. The shared electrons are more likely to be closer to the atom with greater electronegativity, and less likely to be near the atom of lower electronegativity. In this case, although the molecule is still electrically neutral (same number of protons as electrons), the distribution of charge is not uniform. This unequal distribution results in regions of partial negative charge near the more electronegative atom, and regions of partial positive charge near the less electronegative atom. Such bonds are termed **polar covalent bonds,** and the molecules polar molecules. When drawing polar molecules, these partial charges are usually symbolized by the lowercase Greek letter delta (δ). The partial charge seen in a polar covalent bond is relatively small—far less than the unit charge of an ion. For biological molecules, we can predict polarity of bonds by knowing the relative electronegativity of a small number of important atoms (table 2.2). Notice that although C and H differ slightly in electronegativity, this small difference is negligible, and C—H bonds are considered nonpolar.

Because of its importance in the chemistry of water, we will explore the nature of polar and nonpolar molecules in the section 2.4. Water (H_2O) is a polar molecule with electrons more concentrated around the oxygen atom.

Chemical reactions alter bonds

The formation and breaking of chemical bonds, which is the essence of chemistry, is termed a *chemical reaction.* All chemical reactions involve the shifting of atoms from one molecule or ionic compound to another, without any change in the number or identity of the atoms. For convenience, we refer to the original molecules before the reaction starts as *reactants,* and the molecules resulting from the chemical reaction as *products.* For example:

$$6H_2O + 6CO_2 \longrightarrow C_6H_{12}O_6 + 6O_2$$
$$\textit{reactants} \longrightarrow \textit{products}$$

You may recognize this reaction as a simplified form of the photosynthesis reaction, in which water and carbon dioxide are combined to produce glucose and oxygen. Most animal life ultimately depends on this reaction, which takes place in plants. (Photosynthetic reactions will be discussed in detail in chapter 8.)

TABLE 2.2	Relative Electronegativities of Some Important Atoms
Atom	**Electronegativity**
O	3.5
N	3.0
C	2.5
H	2.1

The extent to which chemical reactions occur is influenced by three important factors:

1. **Temperature.** Heating the reactants increases the rate of a reaction because the reactants collide with one another more often. (Care must be taken that the temperature is not so high that it destroys the molecules.)
2. **Concentration of reactants and products.** Reactions proceed more quickly when more reactants are available, allowing more frequent collisions. An accumulation of products typically slows the reaction and, in reversible reactions, may speed the reaction in the reverse direction.
3. **Catalysts.** A catalyst is a substance that increases the rate of a reaction. It doesn't alter the reaction's equilibrium between reactants and products, but it does shorten the time needed to reach equilibrium, often dramatically. In living systems, proteins called enzymes catalyze almost every chemical reaction.

Many reactions in nature are reversible. This means that the products may themselves be reactants, allowing the reaction to proceed in reverse. We can write the preceding reaction in the reverse order:

$$C_6H_{12}O_6 + 6O_2 \longrightarrow 6H_2O + 6CO_2$$
$$\textit{reactants} \longrightarrow \textit{products}$$

This reaction is a simplified version of the oxidation of glucose by cellular respiration, in which glucose is broken down into water and carbon dioxide in the presence of oxygen. Virtually all organisms carry out forms of glucose oxidation; details are covered later, in chapter 7.

Learning Outcomes Review 2.3

An ionic bond is an attraction between ions of opposite charge in an ionic compound. A covalent bond is formed when two atoms share one or more pairs of electrons. Complex biological compounds are formed in large part by atoms that can form one or more covalent bonds: C, H, O, and N. A polar covalent bond is formed by unequal sharing of electrons. Nonpolar bonds exhibit equal sharing of electrons.

- ***How is a polar covalent bond different from an ionic bond?***

2.4 Water: A Vital Compound

Learning Outcomes

1. ***Relate how the structure of water leads to hydrogen bonds.***
2. ***Describe water's cohesive and adhesive properties.***

Of all the common molecules, only water exists as a liquid at the relatively low temperatures that prevail on the Earth's surface. Three-fourths of the Earth is covered by liquid water (figure 2.10).

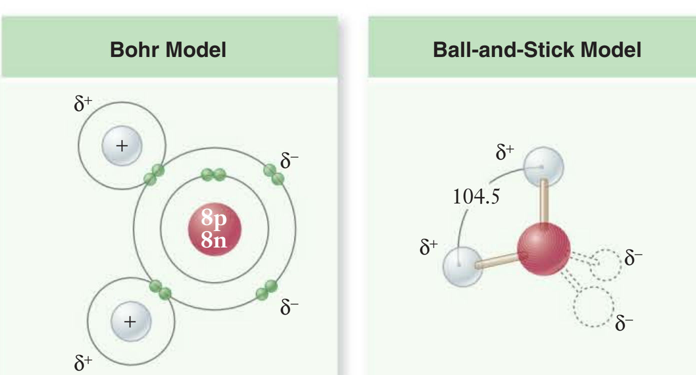
a. Solid *b.* Liquid *c.* Gas

Figure 2.10 Water takes many forms. ***a.*** When water cools below 0°C, it forms beautiful crystals, familiar to us as snow and ice. ***b.*** Ice turns to liquid when the temperature is above 0°C. ***c.*** Liquid water becomes steam when the temperature rises above 100°C, as seen in this hot spring at Yellowstone National Park.

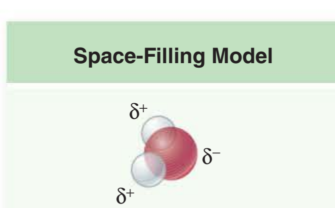
When life was beginning, water provided a medium in which other molecules could move around and interact, without being held in place by strong covalent or ionic bonds. Life evolved in water for 2 billion years before spreading to land. And even today, life is inextricably tied to water. About two-thirds of any organism's body is composed of water, and all organisms require a water-rich environment, either inside or outside it, for growth and reproduction. It is no accident that tropical rain forests are bursting with life, whereas dry deserts appear almost lifeless except when water becomes temporarily plentiful, such as after a rainstorm.

Water's structure facilitates hydrogen bonding

Water has a simple molecular structure, consisting of an oxygen atom bound to two hydrogen atoms by two single covalent bonds (figure 2.11). The resulting molecule is stable: It satisfies the octet rule, has no unpaired electrons, and carries no net electrical charge. The electronegativity of O is much greater than that of H (see table 2.2), and so the bonds between these atoms are highly polar. *The polarity of water underlies water's chemistry and the chemistry of life.*



a. *b.*



c.

Figure 2.11 Water has a simple molecular structure. ***a.*** Each water molecule is composed of one oxygen atom and two hydrogen atoms. The oxygen atom shares one electron with each hydrogen atom. ***b.*** The greater electronegativity of the oxygen atom makes the water molecule polar: Water carries two partial negative charges (δ^-) near the oxygen atom and two partial positive charges (δ^+), one on each hydrogen atom. ***c.*** Space-filling model shows what the molecule would look like if it were visible.

The single most outstanding chemical property of water is its ability to form weak chemical associations, called **hydrogen bonds.** These bonds form between the partially negative O atoms and the partially positive H atoms of two water molecules. Although these bonds have only 5–10% of the strength of covalent bonds, they are important to DNA and protein structure, and thus responsible for much of the chemical organization of living systems.

If we consider the shape of a water molecule, we see that its two covalent bonds have a partial charge at each end: δ^- at the oxygen end and δ^+ at the hydrogen end. The most stable arrangement of these charges is a *tetrahedron (a pyramid with a triangle as its base),* in which the two negative and two positive charges are approximately equidistant from one another. The oxygen atom lies at the center of the tetrahedron, the hydrogen atoms occupy two of the apexes (corners), and the partial negative charges occupy the other two apexes (figure 2.11*b*). The bond angle between the two covalent oxygen–hydrogen bonds is 104.5°. This value is slightly less than the bond angle of a regular tetrahedron, which would be 109.5°. In water, the partial negative charges occupy more space than the partial positive regions, so the oxygen–hydrogen bond angle is slightly compressed.

Water molecules are cohesive

The polarity of water allows water molecules to be attracted to one another—that is, water is *cohesive.* The oxygen end of each water molecule, which is δ^-, is attracted to the hydrogen end, which is δ^+, of other molecules. The attraction produces hydrogen bonds among water molecules (figure 2.12). Each hydrogen bond is individually very weak and transient, lasting on average only a hundred-billionth (10^{-11}) of a second. The cumulative effects of large numbers of these bonds, however, can be enormous. Water forms an abundance of hydrogen bonds, which are responsible for many of its important physical properties (table 2.3).

Water's cohesion is responsible for its being a liquid, not a gas, at moderate temperatures. The cohesion of liquid water is also responsible for its **surface tension.** Small insects can walk on water (figure 2.13) because at the air–water interface, all the surface water molecules are hydrogen-bonded to molecules below them.

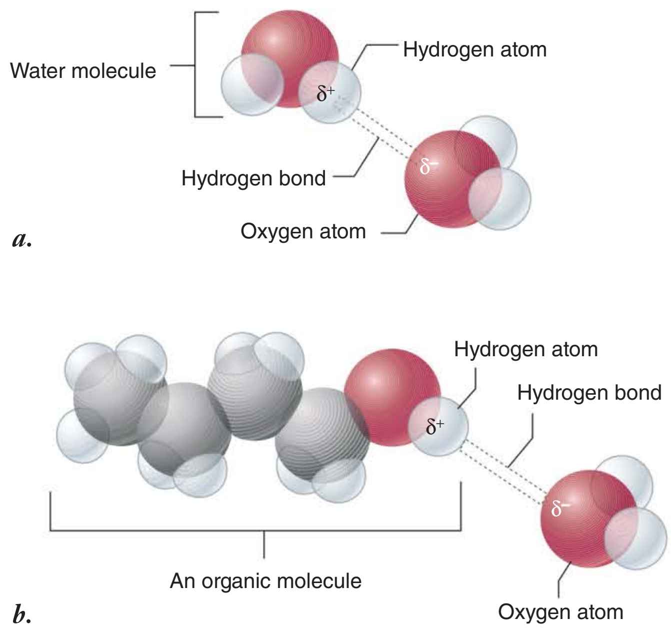


Figure 2.12 Structure of a hydrogen bond. ***a.*** Hydrogen bond between two water molecules. ***b.*** Hydrogen bond between an organic molecule (*n*-butanol) and water. H in *n*-butanol forms a hydrogen bond with oxygen in water. This kind of hydrogen bond is possible any time H is bound to a more electronegative atom (see table 2.2).

Figure 2.13 Cohesion. Some insects, such as this water strider, literally walk on water. Because the surface tension of the water is greater than the force of one foot, the strider glides atop the surface of the water rather than sinking. The high surface tension of water is due to hydrogen bonding between water molecules.

Water molecules are adhesive

The polarity of water causes it to be attracted to other polar molecules as well. This attraction for other polar substances is called *adhesion.* Water adheres to any substance with which it can form hydrogen bonds. This property explains why substances containing polar molecules get "wet" when they are immersed in water, but those that are composed of nonpolar molecules (such as oils) do not.

The attraction of water to substances that have electrical charges on their surface is responsible for capillary action. If a glass tube with a narrow diameter is lowered into a beaker of water, the water will rise in the tube above the level of the water in the beaker, because the adhesion of water to the glass surface, drawing it upward, is stronger than the force of gravity, pulling it downward. The narrower the tube, the greater the electrostatic forces between the water and the glass, and the higher the water rises (figure 2.14).

Figure 2.14 Adhesion. Capillary action causes the water within a narrow tube to rise above the surrounding water level; the adhesion of the water to the glass surface, which draws water upward, is stronger than the force of gravity, which tends to pull it down. The narrower the tube, the greater the surface area available for adhesion for a given volume of water, and the higher the water rises in the tube.

TABLE 2.3 The Properties of Water

Property	Explanation	Example of Benefit to Life
Cohesion	Hydrogen bonds hold water molecules together.	Leaves pull water upward from the roots; seeds swell and germinate.
High specific heat	Hydrogen bonds absorb heat when they break and release heat when they form, minimizing temperature changes.	Water stabilizes the temperature of organisms and the environment.
High heat of vaporization	Many hydrogen bonds must be broken for water to evaporate.	Evaporation of water cools body surfaces.
Lower density of ice	Water molecules in an ice crystal are spaced relatively far apart because of hydrogen bonding.	Because ice is less dense than water, lakes do not freeze solid, allowing fish and other life in lakes to survive the winter.
Solubility	Polar water molecules are attracted to ions and polar compounds, making these compounds soluble.	Many kinds of molecules can move freely in cells, permitting a diverse array of chemical reactions.

Learning Outcomes Review 2.4

Because of its polar covalent bonds, water can form hydrogen bonds with itself and with other polar molecules. Hydrogen bonding is responsible for water's cohesion, the force that holds water molecules together, and its adhesion, which is its ability to "stick" to other polar molecules. Capillary action results from both of these properties.

- ***If water were made of C and H instead of H and O, would it still be cohesive and adhesive?***

2.5 Properties of Water

Learning Outcomes

1. *Illustrate how hydrogen bonding affects the properties of water.*
2. *Explain the relevance of water's unusual properties for living systems.*
3. *Identify the dissociation products of water.*

Water moderates temperature through two properties: its high specific heat and its high heat of vaporization. Water also has the unusual property of being less dense in its solid form, ice, than as a liquid. Water acts as a solvent for polar molecules and exerts an organizing effect on nonpolar molecules. All these properties result from its polar nature.

Water's high specific heat helps maintain temperature

The temperature of any substance is a measure of how rapidly its individual molecules are moving. In the case of water, a large input of thermal energy is required to break the many hydrogen bonds that keep individual water molecules from moving about. Therefore, water is said to have a high **specific heat,** which is defined as the amount of heat 1 g of a substance must absorb or lose to change its temperature by 1 degree Celsius (°C). Specific heat measures the extent to which a substance resists changing its temperature when it absorbs or loses heat. Because polar substances tend to form hydrogen bonds, the more polar it is, the higher is its specific heat. The specific heat of water (1 calorie/g/°C) is twice that of most carbon compounds and nine times that of iron. Only ammonia, which is more polar than water and forms very strong hydrogen bonds, has a higher specific heat than water (1.23 cal/g/°C). Still, only 20% of the hydrogen bonds are broken as water heats from 0° to 100°C.

Because of its high specific heat, water heats up more slowly than almost any other compound and holds its temperature longer. Because organisms have a high water content, water's high specific heat allows them to maintain a relatively constant internal temperature. The heat generated by the chemical reactions inside cells would destroy the cells if not for the absorption of this heat by the water within them.

Water's high heat of vaporization facilitates cooling

The **heat of vaporization** is defined as the amount of energy required to change 1 g of a substance from a liquid to a gas. A considerable amount of heat energy (586 cal) is required to accomplish this change in water. As water changes from a liquid to a gas it requires energy (in the form of heat) to break its many hydrogen bonds. The evaporation of water from a surface cools that surface. Many organisms dispose of excess body heat by evaporative cooling, for example, through sweating in humans and many other vertebrates.

Solid water is less dense than liquid water

At low temperatures, water molecules are locked into a crystal-like lattice of hydrogen bonds, forming solid ice (see figure 2.10*a*). Interestingly, ice is less dense than liquid water because the hydrogen bonds in ice space the water molecules relatively far apart. This unusual feature enables icebergs to float. If water did not have this property, nearly all bodies of water would be ice, with only the shallow surface melting every year. The buoyancy of ice is important ecologically because it means bodies of water freeze from the top down and not the bottom up. Because ice floats on the surface of lakes in the winter and the water beneath the ice remains liquid, fish and other animals keep from freezing.

Polar molecules and ions are soluble in water

Water molecules gather closely around any substance that bears an electrical charge, whether that substance carries a full charge (ion) or a charge separation (polar molecule). For example, sucrose (table sugar) is composed of molecules that contain polar hydroxyl (OH) groups. A sugar crystal dissolves rapidly in water because water molecules can form hydrogen bonds with individual hydroxyl groups of the sucrose molecules. Therefore, sucrose is said to be *soluble* in water. Water is termed the *solvent,* and sugar is called the *solute.* Every time a sucrose molecule dissociates, or breaks away, from a solid sugar crystal, water molecules surround it in a cloud, forming a *hydration shell* that prevents it from associating with other sucrose molecules. Hydration shells also form around ions such as Na^+ and Cl^- (figure 2.15).

Water organizes nonpolar molecules

Water molecules always tend to form the maximum possible number of hydrogen bonds. When nonpolar molecules such as oils, which do not form hydrogen bonds, are placed in water, the water molecules act to exclude them. The nonpolar molecules aggregate, or clump together, thus minimizing their disruption of the hydrogen bonding of water. In effect, they shrink from contact with water, and for this reason they are referred to as **hydrophobic** (Greek *hydros,* "water," and *phobos,* "fearing"). In contrast, polar molecules, which readily form hydrogen bonds with water, are said to be **hydrophilic** ("water-loving").

The tendency of nonpolar molecules to aggregate in water is known as **hydrophobic exclusion.** By forcing the hydrophobic portions of molecules together, water causes these molecules to assume particular shapes. This property can also affect the structure of

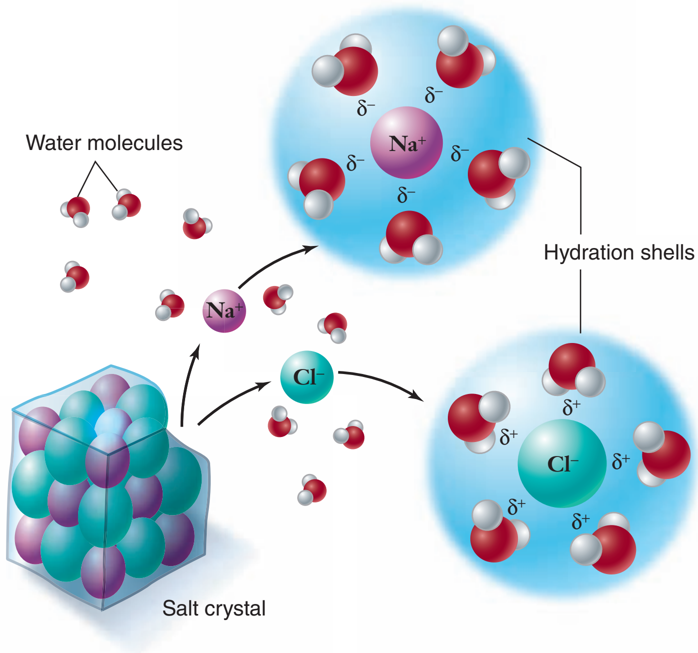


Figure 2.15 Why salt dissolves in water. When a crystal of table salt dissolves in water, individual Na^+ and Cl^- ions break away from the salt lattice and become surrounded by water molecules. Water molecules orient around Na^+ so that their partial negative poles face toward the positive Na^+; water molecules surrounding Cl^- orient in the opposite way, with their partial positive poles facing the negative Cl^-. Surrounded by hydration shells, Na^+ and Cl^- never reenter the salt lattice.

proteins, DNA, and biological membranes. In fact, the interaction of nonpolar molecules and water is critical to living systems.

Water can form ions

The covalent bonds of a water molecule sometimes break spontaneously. In pure water at 25°C, only 1 out of every 550 million water molecules undergoes this process. When it happens, a proton (hydrogen atom nucleus) dissociates from the molecule. Because the dissociated proton lacks the negatively charged electron it was sharing, its positive charge is no longer counterbalanced, and it becomes a hydrogen ion, H^+. The rest of the dissociated water molecule, which has retained the shared electron from the covalent bond, is negatively charged and forms a hydroxide ion, OH^-. This process of spontaneous ion formation is called *ionization:*

$$\underset{\text{water}}{H_2O} \longrightarrow \underset{\text{hydroxide ion}}{OH^-} + \underset{\text{hydrogen ion (proton)}}{H^+}$$

At 25°C, 1 liter (L) of water contains one ten-millionth (or 10^{-7}) mole of H^+ ions. A **mole** (mol) is defined as the weight of a substance in grams that corresponds to the atomic masses of all of the atoms in a molecule of that substance. In the case of H^+, the atomic mass is 1, and a mole of H^+ ions would weigh 1 g. One mole of any substance always contains 6.02×10^{23} molecules of the substance. Therefore, the **molar concentration** of hydrogen ions in pure water, represented as $[H^+]$, is 10^{-7} mol/L. (In reality, the H^+ usually associates with another water molecule to form a hydronium ion, H_3O^+.)

Learning Outcomes Review 2.5

Water has a high specific heat so it does not change temperature rapidly, which helps living systems maintain a near-constant temperature. Water's high heat of vaporization allows cooling by evaporation. Solid water is less dense than liquid water because the hydrogen bonds space the molecules farther apart. Polar molecules are soluble in a water solution, but water tends to exclude nonpolar molecules. Water dissociates to form H^+ and OH^-.

- *How does the fact that ice floats affect life in a lake?*

2.6 Acids and Bases

Learning Outcomes

1. *Define acids, bases, and the pH scale.*
2. *Relate changes in pH to changes in $[H^+]$.*

The concentration of hydrogen ions, and concurrently of hydroxide ions, in a solution is described by the terms *acidity* and *basicity,* respectively. Pure water, having an $[H^+]$ of 10^{-7} mol/L, is considered to be neutral—that is, neither acidic nor basic. Recall that for every H^+ ion formed when water dissociates, an OH^- ion is also formed, meaning that the dissociation of water produces H^+ and OH^- in equal amounts.

The pH scale measures hydrogen ion concentration

The *pH scale* (figure 2.16) is a more convenient way to express the hydrogen ion concentration of a solution. This scale defines *pH,* which stands for "power of hydrogen," as the negative logarithm of the hydrogen ion concentration in the solution:

$$pH = -\log [H^+]$$

Because the logarithm of the hydrogen ion concentration is simply the exponent of the molar concentration of H^+, the pH equals the exponent times –1. For water, therefore, an $[H^+]$ of 10^{-7} mol/L corresponds to a pH value of 7. This is the neutral point—a balance between H^+ and OH^-—on the pH scale. This balance occurs because the dissociation of water produces equal amounts of H^+ and OH^-.

Note that, because the pH scale is *logarithmic,* a difference of 1 on the scale represents a 10-fold change in $[H^+]$. A solution with a pH of 4 therefore has 10 times the $[H^+]$ of a solution with a pH of 5 and 100 times the $[H^+]$ of a solution with a pH of 6.

Acids

Any substance that dissociates in water to increase the $[H^+]$ (and lower the pH) is called an **acid.** The stronger an acid is, the more hydrogen ions it produces and the lower its pH. For example, hydrochloric acid (HCl), which is abundant in your stomach, ionizes completely in water. A dilution of 10^{-1} mol/L of HCl dissociates to form 10^{-1} mol/L of H^+, giving the solution a pH of 1. The pH of champagne, which bubbles because of the carbonic acid dissolved in it, is about 2.

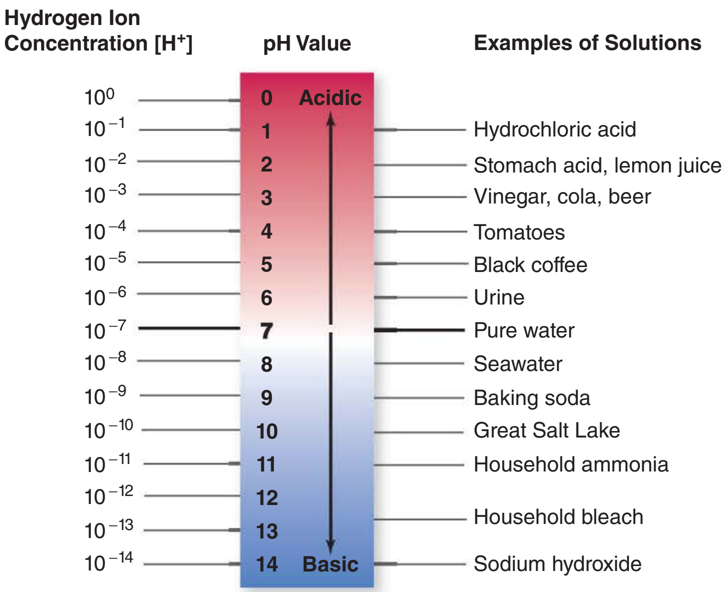


Figure 2.16 The pH scale. The pH value of a solution indicates its concentration of hydrogen ions. Solutions with a pH less than 7 are acidic, whereas those with a pH greater than 7 are basic. The scale is logarithmic, which means that a pH change of 1 represents a 10-fold change in the concentration of hydrogen ions. Thus, lemon juice is 100 times more acidic than tomato juice, and seawater is 10 times more basic than pure water, which has a pH of 7.

Bases

A substance that combines with H^+ when dissolved in water, and thus lowers the $[H^+]$, is called a **base.** Therefore, basic (or alkaline) solutions have pH values above 7. Very strong bases, such as sodium hydroxide (NaOH), have pH values of 12 or more. Many common cleaning substances, such as ammonia and bleach, accomplish their action because of their high pH.

Buffers help stabilize pH

The pH inside almost all living cells, and in the fluid surrounding cells in multicellular organisms, is fairly close to neutral, 7. Most of the enzymes in living systems are extremely sensitive to pH. Often even a small change in pH will alter their shape, thereby disrupting their activities. For this reason, it is important that a cell maintain a constant pH level.

But the chemical reactions of life constantly produce acids and bases within cells. Furthermore, many animals eat substances that are acidic or basic. Cola drinks, for example, are moderately strong (although dilute) acidic solutions. Despite such variations in the concentrations of H^+ and OH^-, the pH of an organism is kept at a relatively constant level by buffers (figure 2.17).

A **buffer** is a substance that resists changes in pH. Buffers act by releasing hydrogen ions when a base is added and absorbing hydrogen ions when acid is added, with the overall effect of keeping $[H^+]$ relatively constant.

Within organisms, most buffers consist of pairs of substances, one an acid and the other a base. The key buffer in human blood is an acid–base pair consisting of carbonic acid (acid) and bicarbonate (base). These two substances interact in a pair of reversible reactions. First, carbon dioxide (CO_2) and H_2O join to form carbonic acid (H_2CO_3), which in a second reaction dissociates to yield bicarbonate ion (HCO_3^-) and H^+.

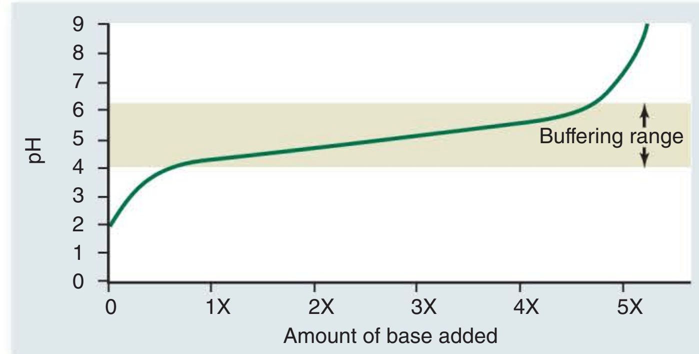
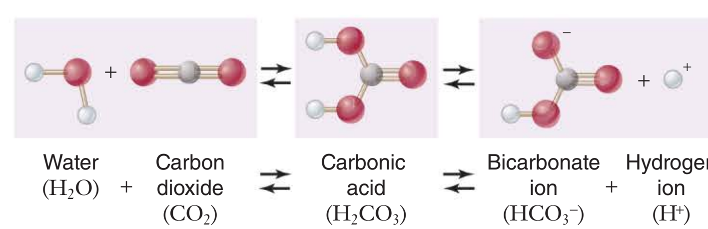


Figure 2.17 Buffers minimize changes in pH. Adding a base to a solution neutralizes some of the acid present, and so raises the pH. Thus, as the curve moves to the right, reflecting more and more base, it also rises to higher pH values. A buffer makes the curve rise or fall very slowly over a portion of the pH scale, called the "buffering range" of that buffer.

Data analysis If we call each step on the *x*-axis one volume of base, how many volumes of base must be added to change the pH from 4 to 6?

If some acid or other substance adds H^+ to the blood, the HCO_3^- acts as a base and removes the excess H^+ by forming H_2CO_3. Similarly, if a basic substance removes H^+ from the blood, H_2CO_3 dissociates, releasing more H^+ into the blood. The forward and reverse reactions that interconvert H_2CO_3 and HCO_3^- thus stabilize the blood's pH:



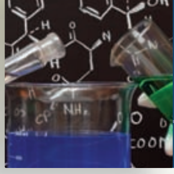
The reaction of carbon dioxide and water to form carbonic acid is a crucial one because it permits carbon, essential to life, to enter water from the air. The Earth's oceans are rich in carbon because of the reaction of carbon dioxide with water.

In a condition called blood acidosis, human blood, which normally has a pH of about 7.4, drops to a pH of about 7.1. This condition is fatal if not treated immediately. The reverse condition, blood alkalosis, involves an increase in blood pH of a similar magnitude and is just as serious.

Learning Outcomes Review 2.6

Acid solutions have a high $[H^+]$, and basic solutions have a low $[H^+]$ (and therefore a high $[OH^-]$). The pH of a solution is the negative logarithm of its $[H^+]$. Low pH values indicate acids, and high pH values indicate bases. Even small changes in pH can be harmful to life. Buffer systems in organisms help to maintain pH within a narrow range.

- *A change of 2 pH units indicates what change in $[H^+]$?*



Chapter Review

2.1 The Nature of Atoms

All matter is composed of atoms (figure 2.3).

Atomic structure includes a central nucleus and orbiting electrons.

Electrically neutral atoms have the same number of protons as electrons. Atoms that gain or lose electrons are called ions.

Elements are defined by the number of protons in the nucleus, the atomic number. Atomic mass is the sum of the mass of protons and neutrons. Isotopes are forms of a single element with different atomic mass due to different numbers of neutrons. Radioactive isotopes are unstable.

Electrons determine the chemical behavior of atoms.

The potential energy of electrons increases as distance from the nucleus increases. Electron orbitals are probability distributions. *s*-Orbitals are spherical; other orbitals have different shapes, such as the dumbbell-shaped *p*-orbitals.

Atoms contain discrete energy levels.

Energy levels correspond to quanta (singular, quantum) of energy, a "ladder" of energy levels that an electron may have.

The loss of electrons from an atom is called oxidation. The gain of electrons is called reduction. Electrons can be transferred from one atom to another in coupled redox reactions.

2.2 Elements Found in Living Systems

The periodic table displays elements according to atomic number and properties.

Atoms tend to establish completely full outer energy levels (the octet rule). Elements with filled outermost orbitals are inert.

Ninety elements occur naturally in the Earth's crust. Twelve of these elements are found in living organisms in greater than trace amounts: C, H, O, N, P, S, Na, K, Ca, Mg, Fe, and Cl.

Compounds of carbon are called organic compounds. The majority of molecules in living systems are composed of C bound to H, O, and N.

2.3 The Nature of Chemical Bonds

Molecules contain two or more atoms joined by chemical bonds. Compounds contain two or more different elements.

Ionic bonds form crystals.

Ions with opposite electrical charges form ionic bonds, such as NaCl (figure 2.9*b*).

Covalent bonds build stable molecules.

A molecule formed by a covalent bond is stable because it has no net charge, the octet rule is satisfied, and it has no unpaired electrons. Covalent bonds may be single, double, or triple, depending on the number of pairs of electrons shared. Nonpolar covalent bonds involve equal sharing of electrons between atoms. Polar covalent bonds involve unequal sharing of electrons.

Chemical reactions alter bonds.

Temperature, reactant concentration, and the presence of catalysts affect reaction rates. Most biological reactions are reversible, such as the conversion of carbon dioxide and water into carbohydrates.

2.4 Water: A Vital Compound

Water's structure facilitates hydrogen bonding.

Hydrogen bonds are weak interactions between a partially positive H in one molecule and a partially negative O in another molecule (figure 2.11).

Water molecules are cohesive.

Cohesion is the tendency of water molecules to adhere to one another due to hydrogen bonding. The cohesion of water is responsible for its surface tension.

Water molecules are adhesive.

Adhesion occurs when water molecules adhere to other polar molecules. Capillary action results from water's adhesion to the sides of narrow tubes, combined with its cohesion.

2.5 Properties of Water

Water's high specific heat helps maintain temperature.

The specific heat of water is high because it takes a considerable amount of energy to disrupt hydrogen bonds.

Water's high heat of vaporization facilitates cooling.

Breaking hydrogen bonds to turn liquid water into vapor takes a lot of energy. Many organisms lose excess heat through evaporative cooling, such as sweating.

Solid water is less dense than liquid water.

Hydrogen bonds are spaced farther apart in the solid phase of water than in the liquid phase. As a result, ice floats.

Polar molecules and ions are soluble in water.

Water's polarity makes it a good solvent for polar substances and ions. Polar molecules or portions of molecules are attracted to water (hydrophilic). Molecules that are nonpolar are repelled by water (hydrophobic). Water makes nonpolar molecules clump together.

Water organizes nonpolar molecules.

Nonpolar molecules will aggregate to avoid water. This maximizes the hydrogen bonds that water can make. This hydrophobic exclusion can affect the structure of DNA, proteins, and biological membranes.

Water can form ions.

Water dissociates into H^+ and OH^-. The concentration of H^+, shown as $[H^+]$, in pure water is 10^{-7} mol/L.

2.6 Acids and Bases (figure 2.16)

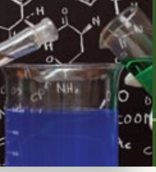
The pH scale measures hydrogen ion concentration.

pH is defined as the negative logarithm of $[H^+]$. Pure water has a pH of 7. A difference of 1 pH unit means a 10-fold change in $[H^+]$.

Acids have a greater $[H^+]$ and therefore a lower pH; bases have a lower $[H^+]$ and therefore a higher pH.

Buffers help stabilize pH.

Carbon dioxide and water react reversibly to form carbonic acid. A buffer resists changes in pH by absorbing or releasing H^+. The key buffer in the human blood is the carbonic acid/bicarbonate pair.



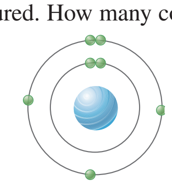
Review Questions

UNDERSTAND

1. The property that distinguishes an atom of one element (carbon, for example) from an atom of another element (oxygen, for example) is
   - a. the number of electrons.
   - b. the number of protons.
   - c. the number of neutrons.
   - d. the combined number of protons and neutrons.
2. If an atom has one valence electron—that is, a single electron in its outer energy level—it will most likely form
   - a. one polar, covalent bond.
   - b. two nonpolar, covalent bonds.
   - c. two covalent bonds.
   - d. an ionic bond.
3. An atom with a net positive charge must have more
   - a. protons than neutrons.
   - b. protons than electrons.
   - c. electrons than neutrons.
   - d. electrons than protons.
4. The isotopes carbon-12 and carbon-14 differ in
   - a. the number of neutrons.
   - b. the number of protons.
   - c. the number of electrons.
   - d. Both b and c are correct.
5. Which of the following is NOT a property of the elements most commonly found in living organisms?
   - a. The elements have a low atomic mass.
   - b. The elements have an atomic number less than 21.
   - c. The elements possess eight electrons in their outer energy level.
   - d. The elements are lacking one or more electrons from their outer energy level.
6. Ionic bonds arise from
   - a. shared valence electrons.
   - b. attractions between valence electrons.
   - c. charge attractions between valence electrons.
   - d. attractions between ions of opposite charge.
7. A solution with a high concentration of hydrogen ions
   - a. is called a base.
   - b. is called an acid.
   - c. has a high pH.
   - d. Both b and c are correct.

APPLY

1. Using the periodic table on page 22, which of the following atoms would you predict should form a positively charged ion (cation)?
   - a. Fluorine (F)
   - b. Neon (Ne)
   - c. Potassium (K)
   - d. Sulfur (S)
2. Refer to the element pictured. How many covalent bonds could this atom form?
   - a. Two
   - b. Three
   - c. Four
   - d. None



3. A molecule with polar covalent bonds would
   - a. be soluble in water.
   - b. not be soluble in water.
   - c. contain atoms with very similar electronegativity.
   - d. Both b and c are correct.
4. Hydrogen bonds are formed
   - a. between any molecules that contain hydrogen.
   - b. only between water molecules.
   - c. when hydrogen is part of a polar bond.
   - d. when two atoms of hydrogen share an electron.
5. If you shake a bottle of oil and vinegar then let it sit, it will separate into two phases because
   - a. the nonpolar oil is soluble in water.
   - b. water can form hydrogen bonds with the oil.
   - c. polar oil is not soluble in water.
   - d. nonpolar oil is not soluble in water.
6. The decay of radioactive isotopes involves changes to the nucleus of atoms. Explain how this differs from the changes in atoms that occur during chemical reactions.

SYNTHESIZE

1. Elements that form ions are important for a range of biological processes. You have learned something about the cations sodium (Na^{+}), calcium ($Ca2^{+}$), and potassium (K^{+}) in this chapter. Use your knowledge of the definition of a cation to identify other examples from the periodic table.

2. A popular theme in science fiction literature has been the idea of silicon-based life-forms in contrast to our carbon-based life. Evaluate the possibility of silicon-based life based on the chemical structure and potential for chemical bonding of a silicon atom.

3. Efforts by NASA to search for signs of life on Mars have focused on the search for evidence of liquid water rather than looking directly for biological organisms (living or fossilized). Use your knowledge of the influence of water on life on Earth to construct an argument justifying this approach.