2 LIFE'S CHEMICAL BASIS

What Are You Worth?

Hollywood thinks Leonardo DiCaprio is worth $20 million per movie, the Yankees think shortstop Alex Rodriguez is worth $217 million per decade, and the United States thinks the average teacher is worth $44,367 per year. Chemically, though, how much is the human body really worth (Figure 2.1*a*)?

Each of us is a collection of **elements**, or fundamental substances that each consist of only one kind of atom. An **atom** is the smallest unit of an element that still retains the element's properties. It occupies space, has mass, and cannot be broken down into something else, at least by everyday means.

Oxygen, carbon, hydrogen, nitrogen, and calcium are the main elements in organisms. Next are phosphorus, potassium, sulfur, sodium, and chlorine. There are a lot of *trace* elements, each making up less than 0.01 percent of the body's weight. Selenium and lead are examples.

Wait a minute! Selenium, lead, mercury, arsenic, and many other elements are toxic, right? So how can they be part of the collection? We're finding that trace amounts

a

Mass of Elements in a 70-Kilogram Human Body		Cost (Retail)
Oxygen	43.00 kilograms (kg)	$0.021739
Carbon	16.00 kg	6.400000
Hydrogen	7.00 kg	0.028315
Nitrogen	1.80 kg	9.706929
Calcium	1.00 kg	15.500000
Phosphorus	780.00 grams (g)	68.198594
Potassium	140.00 g	4.098737
Sulfur	140.00 g	0.011623
Sodium	100.00 g	2.287748
Chlorine	95.00 g	1.409496
Magnesium	19.00 g	0.444909
Iron	4.20 g	0.054600
Fluorine	2.60 g	7.917263
Zinc	2.30 g	0.088090
Silicon	1.00 g	0.370000
Rubidium	0.68 g	1.087153
Strontium	0.32 g	0.177237
Bromine	0.26 g	0.012858
Lead	0.12 g	0.003960
Copper	72.00 milligrams (mg)	0.012961
Aluminum	60.00 mg	0.246804
Cadmium	50.00 mg	0.010136
Cerium	40.00 mg	0.043120
Barium	22.00 mg	0.028776
Iodine	20.00 mg	0.094184
Tin	20.00 mg	0.005387
Titanium	20.00 mg	0.010920
Boron	18.00 mg	0.002172
Nickel	15.00 mg	0.031320
Selenium	15.00 mg	0.037949
Chromium	14.00 mg	0.003402
Manganese	12.00 mg	0.001526
Arsenic	7.00 mg	0.023576
Lithium	7.00 mg	0.024233
Cesium	6.00 mg	0.000016
Mercury	6.00 mg	0.004718
Germanium	5.00 mg	0.130435
Molybdenum	5.00 mg	0.001260
Cobalt	3.00 mg	0.001509
Antimony	2.00 mg	0.000243
Silver	2.00 mg	0.013600
Niobium	1.50 mg	0.000624
Zirconium	1.00 mg	0.000830
Lanthanium	0.80 mg	0.000566
Gallium	0.70 mg	0.003367
Tellurium	0.70 mg	0.000722
Yttrium	0.60 mg	0.005232
Bismuth	0.50 mg	0.000119
Thallium	0.50 mg	0.000894
Indium	0.40 mg	0.000600
Gold	0.20 mg	0.001975
Scandium	0.20 mg	0.058160
Tantalum	0.20 mg	0.001631
Vanadium	0.11 mg	0.000322
Thorium	0.10 mg	0.004948
Uranium	0.10 mg	0.000103
Samarium	50.00 micrograms (µg)	0.000118
Beryllium	36.00 µg	0.000218
Tungsten	20.00 µg	0.000007
Grand Total		$118.63

Figure 2.1 (**a**) What are you worth, chemically speaking? (**b**) Proportions of the most common elements in a human body, Earth's crust, and seawater. How are they similar? How do they differ?

IMPACTS, ISSUES

Human	
Oxygen	61.0%
Carbon	23.0
Hydrogen	10.0
Nitrogen	2.6
Calcium	1.4
Phosphorus	1.1
Potassium	0.2
Sulfur	0.2

Earth's Crust	
Oxygen	46.0%
Silicon	27.0
Aluminum	8.2
Iron	6.3
Calcium	5.0
Magnesium	2.9
Sodium	2.3
Potassium	1.5

Seawater	
Oxygen	85.7%
Hydrogen	10.8
Chlorine	2.0
Sodium	1.1
Magnesium	0.1
Sulfur	0.1
Calcium	0.04
Potassium	0.03

b

of at least some of them have vital functions. For instance, even a little selenium is toxic, but *too* little can cause heart problems and thyroid disorders.

Superficially, then, the human body can be viewed as a balanced collection of elements. The amounts are worth no more than $118.63, and the kinds are not even unique; they occur in Earth's crust and even seawater (Figure 2.1*b*). However, the *proportions* of elements in humans and other organisms are unique relative to nonliving things. Look at all of that carbon, for instance! Also, you will never find a clod of dirt or a volume of seawater that comes close to the *structural and functional organization* of a living body. Assembling that collection of elements into an organized, operational body takes a fabulous molecular library (DNA), enzymes and other metabolic workers, and large, ongoing inputs of energy (just ask any pregnant woman).

Remember this when someone tries to say "chemistry" has nothing to do with you. It has everything to do with you. People, toothpaste, turkeys, refrigerators, jet fuel, health, disease, corsages, acid rain, nerve gas, old-growth forests—name any living or nonliving bit of the universe, and chemistry is part of it.

Watch the video online!

How Would You Vote?

Fluoride helps prevent tooth decay. But too much wrecks bones and teeth, and causes birth defects. A lot can kill you. Many communities in the United States add fluoride to their supply of drinking water. Do you want it in yours? See BiologyNow for details, then vote online.

Key Concepts

ATOMS AND ELEMENTS

An element is a fundamental substance made of one type of atom. The atom is the smallest unit of an element that still retains the element's properties, and its building blocks are protons, electrons, and neutrons. Isotopes are atoms of an element that vary in the number of neutrons. Sections 2.1, 2.2

WHY ELECTRONS MATTER

Atoms acquire, share, and give up electrons. Whether one atom will bond with others depends on the number and arrangement of its electrons. Section 2.3

ATOMS BOND

The bonding behavior of biological molecules starts with the number and arrangement of electrons in each type of atom. Ionic, covalent, and hydrogen bonds are the main categories of bonds between atoms in biological molecules. Section 2.4

NO WATER, NO LIFE

Life originated in water and is adapted to its properties. Water has temperature-stabilizing effects. Many kinds of substances dissolve easily in it. Water also shows cohesion. Section 2.5

HYDROGEN IONS RULE

Life depends on precise controls over the formation, use, and buffering of hydrogen ions. Section 2.6

Links to Earlier Concepts

With this chapter, we start at the base of life's levels of organization, so take a moment to review the simple chart in Section 1.1. It all starts with atoms and energy. Life's organization requires tapping into a great one-way flow of energy and storing it in bonds between atoms (1.2).

The chapter also has a simple example of how the body's built-in mechanisms help return the internal environment to a homeostatic state when conditions shift beyond ranges that cells can tolerate (1.2).

2.1 Start With Atoms

LINK TO SECTION 1.1

Know a bit about protons, neutrons, and electrons, and you have a clue to why the elements that make up the body behave as they do. Each element's unique properties start with the number of protons in its atoms.

An element, again, is a fundamental substance made of only one kind of atom. Atoms are built from three kinds of subatomic particles: protons, electrons, and neutrons. Each **proton** carries a positive *charge*, which is a defined amount of electricity. You can symbolize a proton as p^+. An atom's nucleus, or core region, holds one or more protons. Except for the hydrogen atom, it also holds **neutrons**, which carry no charge. Moving around the atomic nucleus are one or more **electrons**, which carry a negative charge (e^-). Figure 2.2 shows a few simple models for atomic structure.

The positive charge of one proton and the negative charge of one electron balance each other. Therefore, an atom that has the same number of electrons and protons has no net electrical charge.

Each element has a unique *atomic number*, which is the number of protons in the nucleus of its atoms. A hydrogen atom has one proton, so the atomic number is 1. For carbon, with six protons, it is 6.

Protons and neutrons contribute to an atom's mass. (Electrons are too tiny to do so.) We can assign each element a *mass number*, or the total number of protons *and* neutrons in the atomic nucleus. For carbon, with six protons and six neutrons, the mass number is 12.

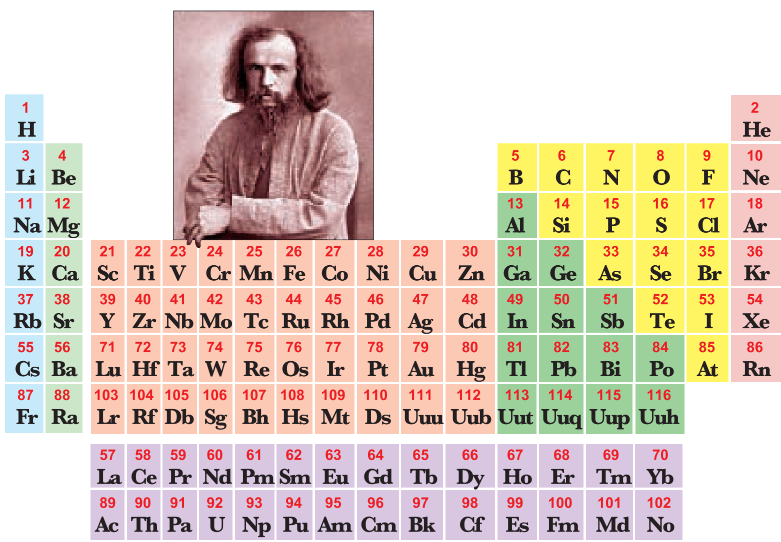


Figure 2.3 Periodic table of the elements and Dmitry Mendeleev, who created it. Some symbols for elements are abbreviations for their Latin names. For instance, Pb (lead) is short for *plumbum;* the word "plumbing" is related, because ancient Romans used lead to make their water pipes.

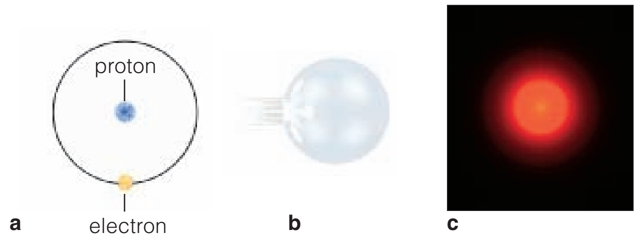


Figure 2.2 Different ways to represent atoms, using hydrogen (H) as the example. (**a**) The shell model, good for showing the number of electrons and their organization around the nucleus. (**b**) Ball models show the sizes of atoms relative to one another. (**c**) Electron density clouds are best at conveying the distribution of electrons around the nucleus.

Why bother with the number of electrons, protons, and neutrons? Knowing them can help you predict how each kind of element will behave under a variety of conditions inside and outside the body.

Elements were being classified in terms of chemical similarities long before their subatomic particles were discovered. In 1869, Dmitry Mendeleev, known more for his extravagant hair than his discoveries (he cut it only once a year), arranged the known elements in a repeating pattern, based on their chemical properties. By using gaps in this **periodic table of the elements**, Mendeleev correctly predicted the existence of many elements that had not yet been discovered.

Elements fall into order in the table according to their atomic number (Figure 2.3). All elements in each vertical column have the same number of electrons that are available for interaction with other atoms. As a result, they behave in similar ways. For example, helium, neon, radon, and other gases in the farthest right column of the periodic table are *inert* elements. Not one of the electrons in their atoms is available for chemical interactions. Such elements rarely do much; they occur mostly as solitary atoms.

You won't find all of the elements in nature. Those after atomic number 92 are extremely unstable. Some have been formed in exceedingly small quantities in laboratories—sometimes no more than a single atom—and they wink out of existence fast.

An element is a fundamental substance consisting of only one kind of atom. Atoms are the smallest units that retain an element's properties.

An atom consists of one or more positively charged protons, negatively charged electrons, and (except for hydrogen) neutrons. Whether any given atom will interact with others depends on how many electrons it has.

2.2 Putting Radioisotopes To Use

FOCUS ON SCIENCE

All elements are defined by the number of protons in their atoms—but an element's atoms can differ in their number of neutrons. We call such atoms ***isotopes*** *of the same element. Some are radioactive.*

In 1896, Henri Becquerel made a chance discovery. He had placed some uranium crystals in a desk drawer, next to a coin and metal screen on top of some sheets of opaque black paper. Underneath that paper was a photographic plate. A day later, the physicist used the film and developed it. Oddly, a negative image of the coin and the metal screen showed up. Becquerel hypothesized that energy radiating from the uranium salts had passed through the paper—which was impenetrable to light—and exposed the film around both metal objects.

As we now know, uranium has isotopes—fifteen of them. Most naturally occurring elements do. Carbon has three isotopes, nitrogen has two, and so on. A superscript number to the left of an element's symbol is the isotope's mass number. For instance, carbon's three natural isotopes are ^{12}C (carbon 12, the most common form, with six protons, six neutrons), ^{13}C (with six protons, seven neutrons), and ^{14}C (with six protons, eight neutrons).

Some isotopes are unstable, or radioactive. A radioactive isotope, or **radioisotope**, spontaneously emits energy in the form of subatomic particles and x-rays when its nucleus disintegrates. This process is called **radioactive decay**, and it can transform one element into another. As an example, ^{13}C and ^{14}C are radioisotopes of carbon. Each predictably decays with a particular amount of energy into a more stable product. After 5,700 years, about half of the atoms in a sample of ^{14}C will have turned into ^{13}N (nitrogen) atoms. Researchers use radioactive decay to estimate the age of rocks and biological remains, as Section 17.5 explains.

The different isotopes of an element are still the same element. For the most part, carbon is carbon, regardless of how many neutrons it has. Living systems use ^{12}C the same way as ^{14}C. Knowing this, researchers or clinicians who want to track a particular substance construct a **tracer**. Tracers are molecules in which a radioisotope has been substituted for a more stable isotope. They can be delivered into a cell or multicelled body, even into populations used in laboratory experiments. The energy from radioactive decay is like a shipping label. It helps researchers track the pathway or destination of a substance of interest with the help of radioactivity-detecting instruments.

For example, Melvin Calvin and his colleagues used a tracer to discover specific reaction steps of photosynthesis. They let growing plants take up a radioactive gas (carbon dioxide made with ^{14}C). By using radioactivity-detecting instruments, they tracked the carbon radioisotope through steps by which plants produce simple sugars and starches.

Radioisotopes also are used in medicine. *PET* (short for Positron-Emission Tomography) uses radioisotopes to study metabolism. Clinicians attach a radioisotope to glucose or another sugar. They inject this tracer into a patient, who is moved into a PET scanner (Figure 2.4*a*). Cells in different parts of the body absorb the tracer at different rates. The scanner detects radiation caused by energy from the decay of the radioisotope. That radiation is used to form an image on a monitor, as in Figure 2.4. Such images reveal variations and abnormalities in metabolic activity.

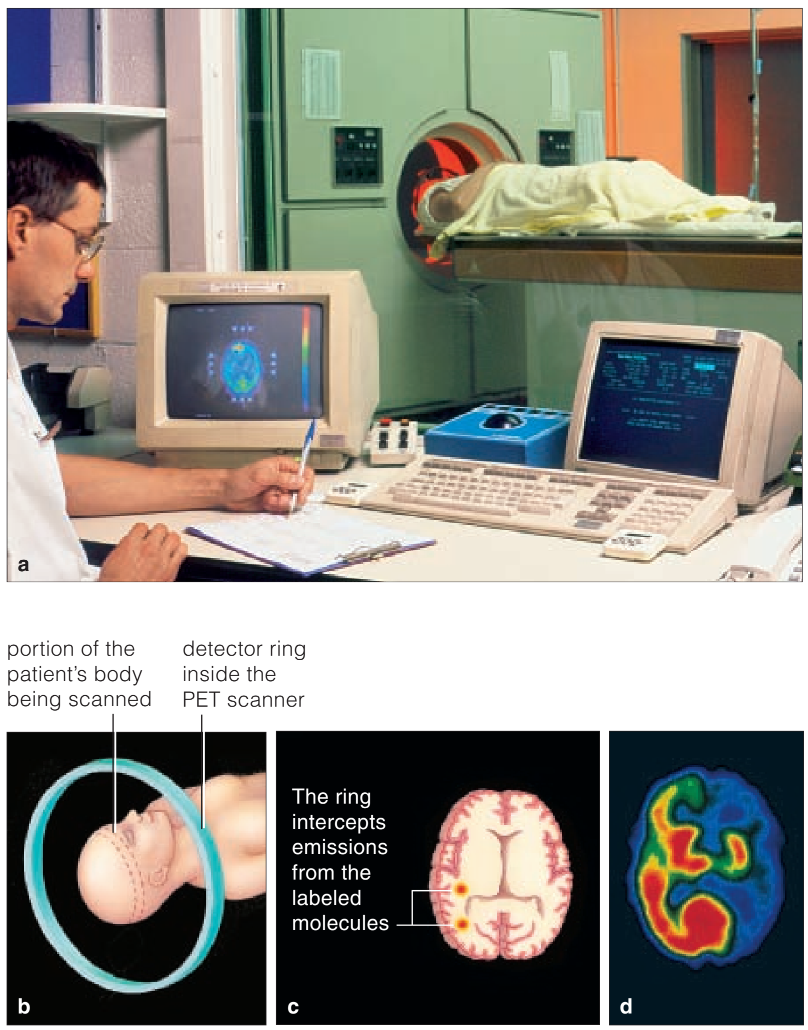


Figure 2.4 ***Animated!*** (**a**) Patient whose brain is being probed in a PET scanner. (**b**,**c**) A ring of detectors intercepts radioactive emissions from tracers that had been injected into the patient. The body region of interest is scanned. Computers analyze and color-code the number of emissions from each location in the scanned region. Results are converted into digital images and displayed on computer screens.

(**d**) Different colors in a scan signify differences in metabolic activity. Cells in the left half of this brain absorbed and used labeled molecules at expected rates. Cells in the right half showed little activity. This patient has a neurological disorder.

2.3 What Happens When Atom Bonds With Atom?

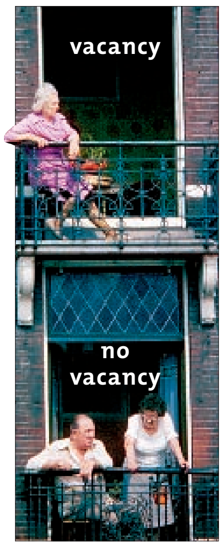
LINK TO SECTION 1.1

Atoms acquire, share, and donate electrons. The atoms of some elements do this quite easily; others do not. Why is this so? To come up with an answer, look to the number and arrangement of electrons in atoms.

ELECTRONS AND ENERGY LEVELS

In our world, simple physics explains the motion of an apple falling from a tree. Tiny electrons belong to a strange world where everyday physics does not apply. (If electrons were as big as apples, you would be 3.5 times taller than our solar system is wide.) Different forces bring about the motion of electrons, which can get from here to there without going in between!

We can calculate where an electron is, although not exactly. The best we can do is say that it is somewhere in a fuzzy cloud of probability density. Where it can go in the cloud depends on how many other electrons belong to the atom. The electrons become arranged in orbitals, or volumes of space around the atomic nucleus. Many orbitals, each with a characteristic three-dimensional shape, are possible.



An atom has the same number of electrons as protons. Most atoms have many electrons. How are they all arranged, given that electrons repel each other? Think of each atom as a multilevel apartment building with lots of vacant rooms to rent to electrons, and a nucleus in the basement. Each "room" is one orbital, and it rents out to two electrons at most. An orbital holding one electron only has a vacancy; another electron can move in.

Each floor in that atomic apartment building corresponds to an energy level. There is only one room on the first floor (one orbital at the lowest energy level, closest to the nucleus). It fills first. For hydrogen, the simplest atom, a lone electron occupies the room (Figure 2.5). Helium, with its two electrons, has no vacancies at the first (lowest) energy level. In larger atoms, more electrons rent the second-floor rooms. If the second floor is filled, then more electrons rent third-floor rooms, and so on. *Electrons fill orbitals at successively higher energy levels.*

The farther an electron is from the basement (the nucleus), the greater its energy. An electron in a first-floor room can't move to the second or third floor, let alone the penthouse, unless a boost of energy puts it there. Suppose the electron absorbs just enough energy from, say, sunlight, to get excited about moving up. Move it does. If nothing fills that lower room, though, the electron will quickly return to it, emitting extra energy as it does. Later on, you will see how cells in plants and in your eyes harness and use that energy.

FROM ATOMS TO MOLECULES

In shell models for atoms, nested "shells" correspond to energy levels. They give us an easy way to check for electron vacancies, as in Figure 2.6. Bear in mind, atoms do not look like these flat diagrams. The shells are not three-dimensional volumes of space, and they certainly don't show the electron orbitals.

The atoms with vacancies in their outermost shell tend to give up, acquire, or share electrons. Actually, what we call **chemical bonds** are just a case of atoms sharing their electrons with one another. An atom with no vacancies rarely bonds with others. But the most common atoms in organisms—such as oxygen, carbon, hydrogen, nitrogen, and calcium—do have vacancies in orbitals at their outermost energy level. They tend to participate in bonds.

A **molecule** is simply two or more atoms of the same or different elements joined in a chemical bond.

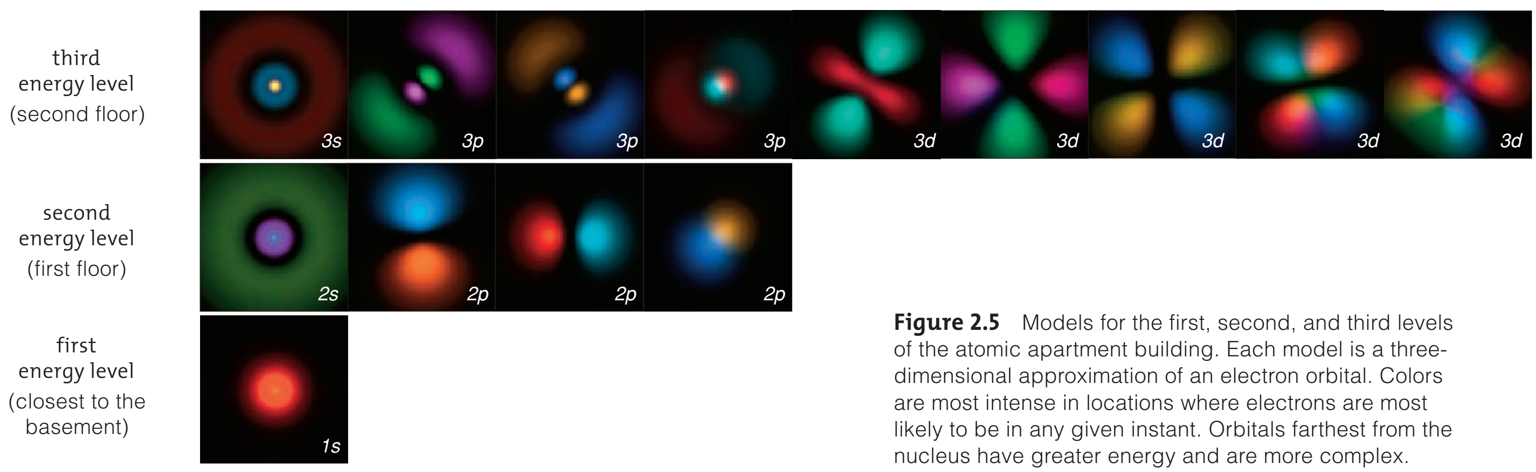


Figure 2.5 Models for the first, second, and third levels of the atomic apartment building. Each model is a three-dimensional approximation of an electron orbital. Colors are most intense in locations where electrons are most likely to be in any given instant. Orbitals farthest from the nucleus have greater energy and are more complex.

c Third shell This shell corresponds to the third energy level. It has nine orbitals (one *s*, three *p*, and five *d* orbitals), or room for eighteen electrons. Sodium has one electron in the third shell of orbitals, and chlorine has seven. Both have vacancies, so they both are receptive to chemical bonding.

b Second shell This shell, which corresponds to the second energy level, has one *s* orbital and three *p* orbitals—room for a total of eight electrons. Carbon has six electrons, two in the first shell and four in the second shell. It has four vacancies. Oxygen has two vacancies. Both carbon and oxygen form chemical bonds.

a First shell A single shell corresponds to the first energy level, which has a single orbital (*1s*) that can hold two electrons. Hydrogen has only one electron in this shell and gives it up easily. A helium atom has two electrons (no vacancies) and usually does not enter into chemical bonds.

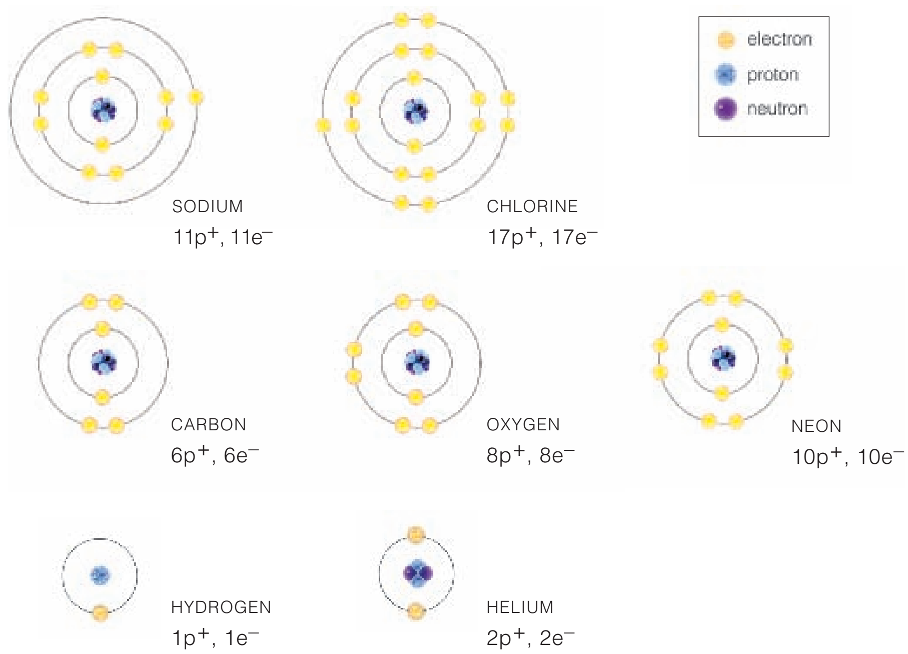


Figure 2.6 ***Animated!*** Shell models, which help us visualize vacancies in an atom's outermost orbitals. Each circle, or shell, represents all orbitals at one energy level. Larger circles correspond to higher energy levels. Such models are highly simplified. A more realistic rendering would show electrons as fuzzy clouds of probability density about 10,000 times larger than the nucleus.

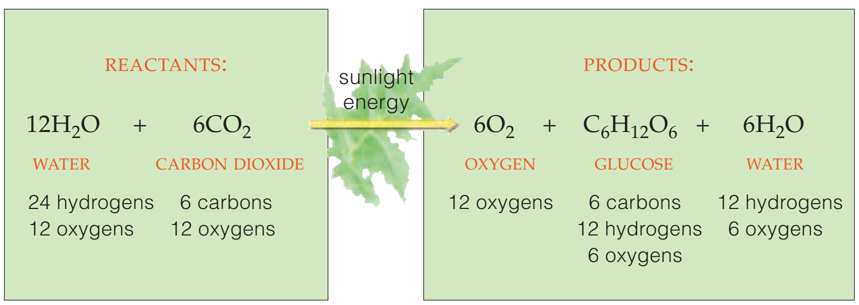


Figure 2.7 Chemical bookkeeping. We use formulas when writing out chemical equations, which represent reactions between atoms and molecules. Substances entering a reaction (reactants) are written to the left of a reaction arrow, and products to the right. How many molecules (or atoms) enter as reactants or form as products are indicated by a number that precedes their formula. *The same number of atoms that enter a reaction must be there at the end.* The atoms get shuffled around, but they never vanish. To be sure you wrote an equation correctly, count the atoms.

You can write a molecule's chemical composition as a formula, which uses symbols for the elements present and subscripts for the number of atoms of each kind of element (Figure 2.7). For example, one molecule of water has the chemical formula H_2O. The subscript number shows that there are two hydrogen (H) atoms for each oxygen (O) atom. If you have six molecules of water, then you would write $6H_2O$.

Compounds are molecules that consist of two or more different elements in proportions that never do vary. Water is an example. All water molecules have one oxygen atom bonded to two hydrogen atoms. The ones in rain clouds, the seas, a Siberian lake, flower petals, your bathtub, or anywhere else have twice as many hydrogen atoms as oxygen atoms. In a **mixture**, two or more substances intermingle without bonding. For example, when you swirl sugar into water, you make a mixture. The proportions of elements in this mixture, or any other kind of mixture, can vary.

Electrons occupy orbitals, or defined volumes of space around an atom's nucleus. Successive orbitals correspond to levels of energy, which become higher with distance from the atomic nucleus.

One or at most two electrons can occupy any orbital. The atoms with vacancies in orbitals at their highest level tend to interact and form bonds with other atoms.

A molecule is two or more atoms joined in a chemical bond. In compounds, atoms of two or more elements are bonded together. A mixture consists of intermingled substances.

ATOMS BOND

2.4 Major Bonds in Biological Molecules

Electrons of one type of atom interact with electrons of others in specific ways. Those interactions give rise to the distinctive properties of biological molecules.

ION FORMATION AND IONIC BONDING

An electron, recall, has a negative charge equal to a proton's positive charge. When an atom contains as many electrons as protons, these charges balance each other, so the atom has a net charge of zero. When an atom *gains* an extra electron, it acquires a net negative charge. When an atom *loses* an electron, it acquires a net positive charge. Either way, it has become an **ion**.

Consider: A chlorine atom has seven protons. It has seven electrons (one vacancy) in the third orbital level—which is most stable when filled with eight. This atom tends to attract an electron from someplace else. With that extra electron, it becomes a chloride ion (Cl^-), with a net negative charge.

Also consider: A sodium atom has eleven protons and eleven electrons. Its second orbital level is full of electrons, and only one electron is in the third orbital level. Giving up the one electron is easier than getting seven more. When it does so, the atom still has eleven protons. But now it has ten electrons. It has become a sodium ion, with a net positive charge (Na^+).

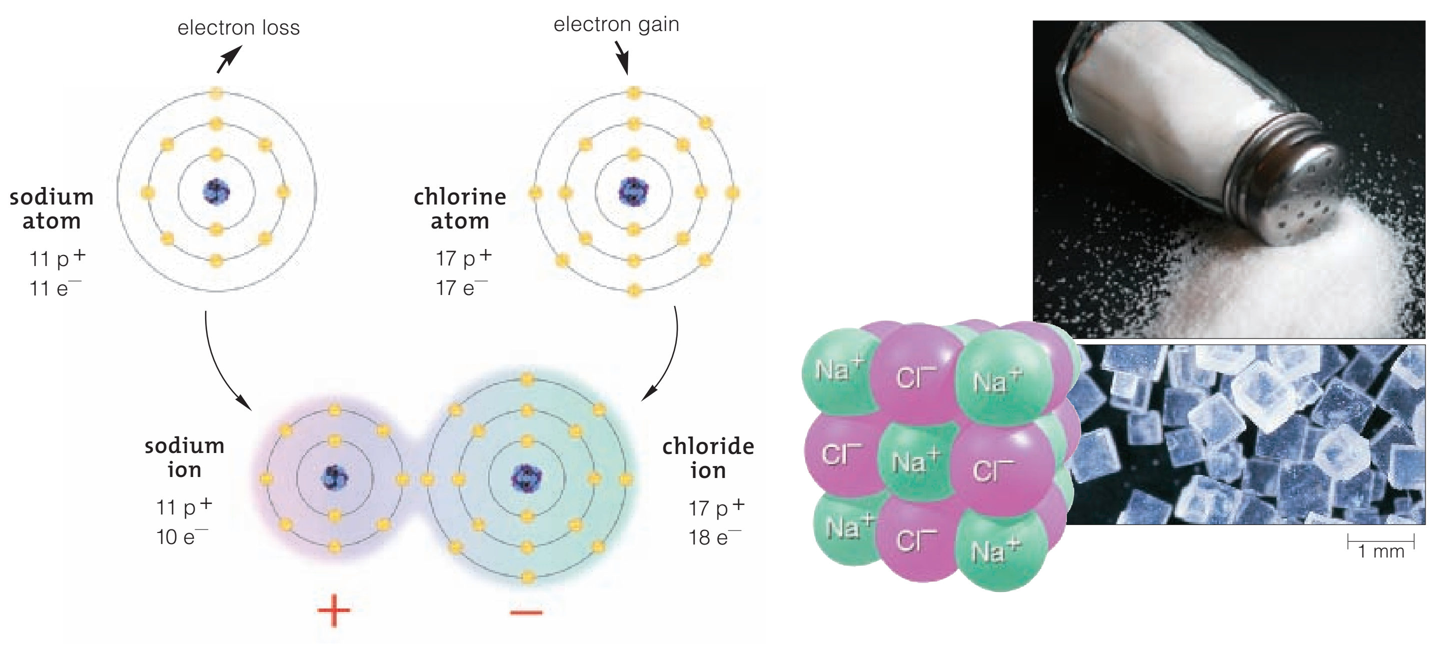
Remember that opposite charges attract each other. When a positively charged ion encounters a negatively charged ion, the two may associate closely with each other. A close association of ions is an **ionic bond**. For example, Figure 2.8*a* shows a crystal of table salt, or NaCl. In such crystals, ionic bonds hold ions of sodium and chloride in an orderly, cubic arrangement.

COVALENT BONDING

In an ionic bond, an atom that has lost one or more electrons associates with an atom that gained one or more electrons. What if both atoms have room for an extra electron? They can *share* one in a hybrid orbital that spans both atomic nuclei. The vacancy in each atom becomes filled with the shared electrons.

When atoms share two electrons, they are joined in a single **covalent bond** (Figure 2.8*b*). Such bonds are stable and are much stronger than ionic bonds.

We can represent covalent bonds as single lines in structural formulas, which show how the atoms of a molecule are physically arranged. A line between two atoms represents a pair of electrons that are being shared in a single covalent bond. To give examples of this bonding pattern, molecular hydrogen (H_2) has one covalent bond and can be written as H—H. Two



a Example of ongoing interactions called ionic bonding. In each crystal of table salt, or NaCl, many sodium ions and chloride ions are staying close together because of the mutual attraction of their opposite charges.

Figure 2.8 ***Animated!*** Important bonds in biological molecules.

atoms share two electron pairs in a *double* covalent bond. Molecular oxygen (O═O) is like this. Others share three electron pairs in a *triple* covalent bond, as in molecular nitrogen (N≡N). Each time you take a breath, O_2 and N_2 molecules flow toward your lungs.

In a *nonpolar* covalent bond, two atoms are sharing electrons equally, so the molecule shows no difference in charge between the two "ends" of the bond. We find such bonds in molecular hydrogen (H_2), oxygen (O_2), and nitrogen (N_2).

In a *polar* covalent bond, two atoms do not share electrons equally. Why not? The atoms are of different elements, and one has more protons than the other. The one with the most protons exerts more of a pull on the electrons, so its end of the bond ends up with a slight negative charge. We say it is "electronegative." The atom at the other end of the bond ends up with a slight positive charge. For instance, a water molecule (H—O—H) has two polar covalent bonds. The oxygen atom carries a slight negative charge, and each of its two hydrogen atoms carries a slight positive charge.

HYDROGEN BONDING

A **hydrogen bond** is a weak attraction that has formed between a covalently bound hydrogen atom and an electronegative atom in a different molecule or in a different region of the same molecule.

Because hydrogen bonds are weak, they form and break easily. Collectively, however, many hydrogen bonds contribute to the properties of liquid water, as you will see next.

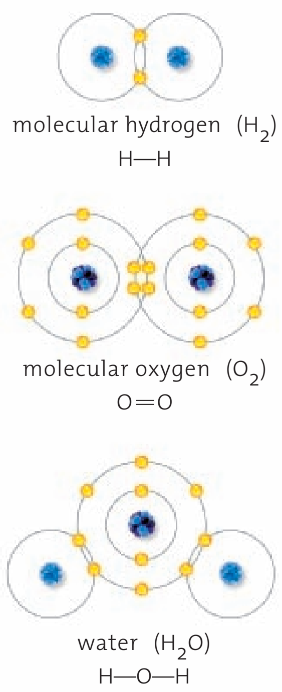
Hydrogen bonds also play important roles in the structure and function of biological molecules. They often form between different parts of large molecules that have folded over on themselves and hold them in particular shapes. Many of these bonds hold DNA's two nucleotide strands together. Figure 2.8*c* hints at the number of these interactions in DNA.

> *Ions form when atoms acquire a net charge by gaining or losing electrons. Two ions of opposite charge attract each other. They can associate in an ionic bond.*
>
> *In a covalent bond, atoms share a pair of electrons. When atoms share the electrons equally, the bond is nonpolar. When the sharing is not equal, the bond is polar—slightly positive at one end, slightly negative at the other.*
>
> *In a hydrogen bond, a covalently bound hydrogen atom attracts a small, negatively charged atom in a different molecule or in a different region of the same molecule.*

Two hydrogen atoms, each with one proton, share two electrons in a single nonpolar covalent bond.

Two oxygen atoms, each with eight protons, share four electrons in a nonpolar double covalent bond.

Oxygen has vacancies for two electrons in its highest energy level orbitals. Two hydrogen atoms can each share an electron with oxygen. The resulting two polar covalent bonds form a water molecule.

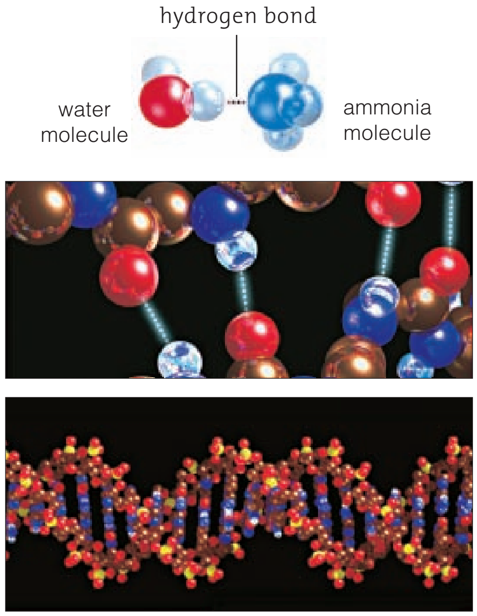


b Covalent bonding. Each atom becomes more stable by sharing electron pairs in hybrid orbitals.

Two molecules interacting weakly in one H bond, which can form and break easily.

H bonds helping to hold part of two large molecules together.

Many H bonds hold DNA's two strands together along their length. Individually each one is weak, but collectively they can stabilize DNA's large structure.



c Hydrogen bonds. Such bonds can form at a hydrogen atom that is already covalently bonded in a molecule. The atom's slight positive charge weakly attracts an atom with a slight negative charge that is already covalently bonded to something else. As shown, this can happen between one of the hydrogen atoms of a water molecule and the nitrogen atom of an ammonia molecule.

2.5 Water's Life-Giving Properties

No sprint through basic chemistry is complete unless it leads to the collection of molecules called water. Life originated in water. Organisms still live in it or they cart water around with them inside cells and tissue spaces. Many metabolic reactions use water. A cell's structure and shape absolutely depend on it.

POLARITY OF THE WATER MOLECULE

Figure 2.9*a* shows the structure of a water molecule. Two atoms of hydrogen have formed polar covalent bonds with an oxygen atom. The molecule has no net charge. Even so, the oxygen pulls the shared electrons more than the hydrogen atoms do. Thus, the molecule of water has a slightly negative "end" that is balanced out by its slightly positive "end."

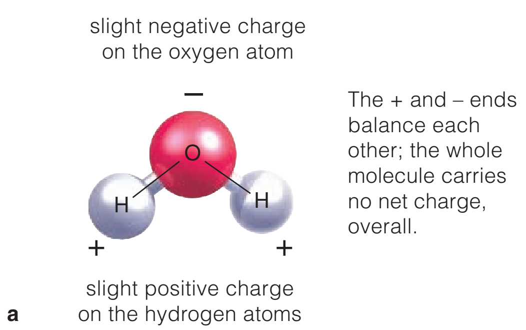
The water molecule's polarity attracts other water molecules. Also, the polarity is so attractive to sugars and other polar molecules that hydrogen bonds form easily between them. That is why polar molecules are known as **hydrophilic** (water-loving) substances.

That same polarity repels oils and other nonpolar molecules, which are **hydrophobic** (water-dreading) substances. Shake a bottle filled with water and salad oil, then set it on a table. Soon, new hydrogen bonds replace the ones that the shaking broke. The reunited water molecules push out oil molecules, which cluster as oil droplets or as an oily film at the water's surface.

The same kinds of interactions proceed at the thin, oily membrane between the water inside and outside cells. Membrane organization—and life itself—starts with such hydrophilic and hydrophobic interactions. You will read about membrane structure in Chapter 5.

WATER'S TEMPERATURE-STABILIZING EFFECTS

Cells are mostly water, and they also release a lot of metabolic heat. The many hydrogen bonds in water keep cells from cooking in their own juices. How? All bonds vibrate nonstop, and they move more as they absorb heat. **Temperature** is a measure of molecular motion. Compared to most other fluids, water absorbs more heat energy before it gets measurably hotter. So water serves as a heat reservoir, and its temperature remains relatively stable. Over time, increases in heat step up the motion within water molecules. Before that happens, however, much of the heat will go into disrupting hydrogen bonds between molecules.



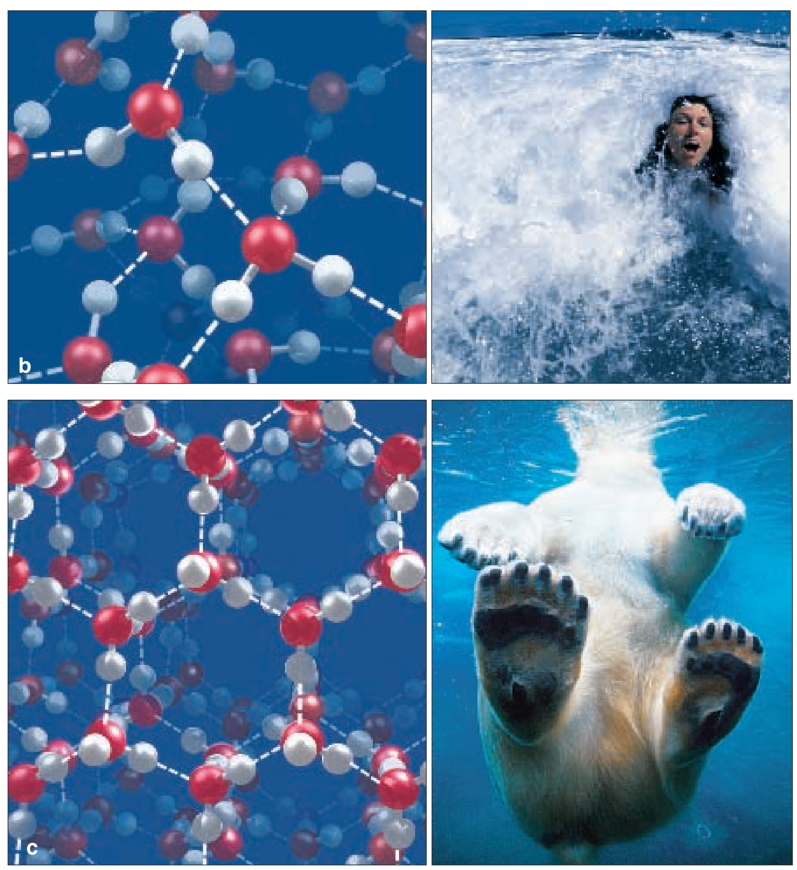


Figure 2.9 *Animated!* Water, a substance essential for life.

(**a**) Polarity of an individual water molecule.

(**b**) Hydrogen bonding pattern among water molecules in liquid water. Dashed lines signify hydrogen bonds, which break and re-form rapidly.

(**c**) Hydrogen bonding in ice. Below 0°C, every water molecule hydrogen-bonds with four others, in a rigid three-dimensional lattice. The molecules are farther apart, or less densely packed, than they are in liquid water. As a result, ice floats on water.

Thanks partly to rising levels of methane and other greenhouse gases that are contributing to global warming, the Arctic ice cap is melting. At current rates, it will be gone in fifty years. So will the polar bears. Already their season for hunting seals is shorter, bears are thinner, and they are giving birth to fewer cubs.

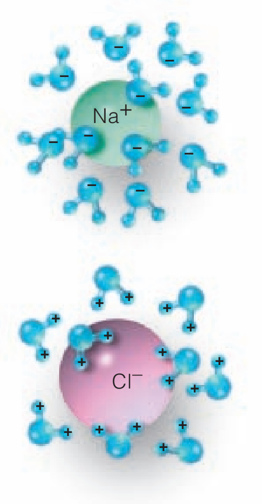
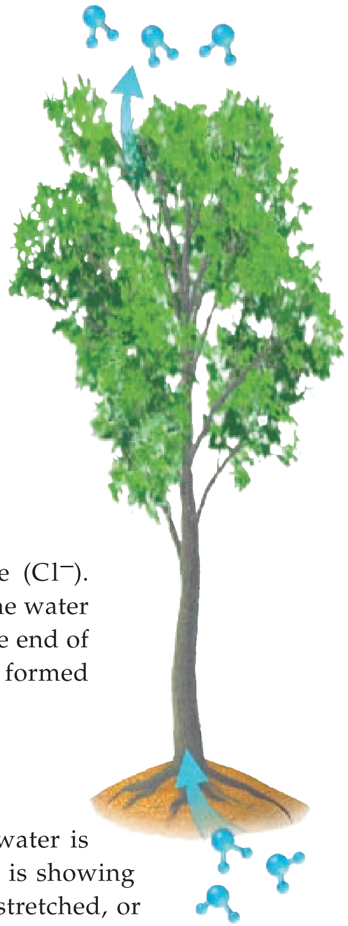


Figure 2.10 Spheres of hydration around two ions.

Figure 2.11 Examples of water's cohesion. (**a**) When a pebble hits liquid water and forces molecules away from the surface, the individual water molecules do not fly every which way. They stay together in droplets. Why? Countless hydrogen bonds exert a continuous inward pull on individual molecules at the surface.

(**b**) And just how does water rise to the very top of trees? Cohesion, and evaporation from leaves, pulls it upward.

a



b

When water temperature is stable, hydrogen bonds form as fast as they break. When water gets hotter, the increase in molecular motion can keep the bonds broken, so individual molecules at the water's surface can escape into air. By this process, **evaporation**, heat energy converts liquid water to gaseous form. The increased energy has overcome the attraction between water molecules, which break free. Water's surface temperature decreases during evaporation.

Evaporative water loss helps you and some other mammals cool off when you sweat on hot, dry days. Sweat, about 99 percent water, evaporates from skin.

Below 0°C, water molecules do not move enough to break their hydrogen bonds, so they become locked in the latticelike bonding pattern of ice (Figure 2.9*c*). Ice is less dense than water. During winter freezes, ice sheets may form near the surface of ponds, lakes, and streams. The ice "blanket" insulates the liquid water beneath it and helps protect many fishes, frogs, and other aquatic organisms against freezing.

WATER'S SOLVENT PROPERTIES

Water is an excellent *solvent*, meaning ions and polar molecules easily dissolve in it. A dissolved substance is known as a **solute**. In general, a substance is said to be *dissolved* after water molecules cluster around ions or molecules of it and keep them dispersed in fluid.

A clustering of water molecules around a solute is called a *sphere of hydration*. Such spheres form around any solute in cellular fluids, tree sap, blood, the fluid in your gut, and every other fluid associated with life. Watch it happen after you pour table salt (NaCl) into a cup of water. In time, the crystals of salt separate into ions of sodium (Na^+) and chloride (Cl^-). Each Na^+ attracts the negative end of some water molecules even as Cl^- attracts the positive end of others (Figure 2.10). Spheres of hydration formed this way keep the ions dispersed in fluid.

WATER'S COHESION

Still another life-sustaining property of water is its cohesion. **Cohesion** means something is showing a capacity to resist rupturing when it is stretched, or placed under tension. You see its effect when a tossed pebble breaks the surface of a lake, a pond, or some other body of liquid water (Figure 2.11*a*). At or near the surface, uncountable numbers of hydrogen bonds are exerting a continuous, inward pull on individual molecules. Bonding creates a high surface tension.

Cohesion is working inside organisms, too. Plants, for example, absorb nutrient-laden water when they grow. Columns of liquid water rise inside pipelines of vascular tissues, which extend from roots to leaves. Water evaporates from leaves when molecules break free and diffuse into air (Figure 2.11*b*). The cohesive force of hydrogen bonds pulls replacements into the leaf cells, in ways explained in Section 30.3.

Being polar, water molecules hydrogen-bond to one another and to other polar (hydrophilic) substances. They tend to repel nonpolar (hydrophobic) substances.

The unique properties of liquid water make life possible. Water has temperature-stabilizing effects, cohesion, and a capacity to dissolve many substances easily.

HYDROGEN IONS RULE

2.6 Acids and Bases

LINK TO SECTION 1.2

Ions dissolved in fluids inside and outside each living cell influence its structure and function. Among the most influential are hydrogen ions. They have far-reaching effects largely because they are chemically active and because there are so many of them.

THE pH SCALE

At any instant in liquid water, some water molecules split into ions of hydrogen (H^+) and hydroxide (OH^-). These ions are the basis of the **pH scale**. The scale is a way to measure the concentration of hydrogen ions in solutions such as seawater, blood, or sap. The greater the H^+ concentration, the lower the pH. Pure water (not rainwater or tap water) always has as many H^+ as OH^- ions. This state is neutrality, or pH 7.0 (Figure 2.12).

A decrease in pH by just one unit from neutrality corresponds to a tenfold increase in H^+ concentration, and an increase by one unit corresponds to a tenfold decrease in H^+ concentration. One way to get a sense of the range is to taste dissolved baking soda (pH 9), water (pH 7), and lemon juice (pH 2).

HOW DO ACIDS AND BASES DIFFER?

When dissolved in water, substances called **acids** *donate* hydrogen ions and **bases** *accept* hydrogen ions. *Acidic* solutions, such as lemon juice, gastric fluid, and coffee, release H^+; their pH is below 7. *Basic* solutions, such as seawater and egg white, easily combine with H^+. Basic solutions, which also are known as alkaline solutions, have a pH above 7.

Nearly all of life's chemistry occurs near pH 7. Most of your body's internal environment (tissue fluids and blood) is between pH 7.3 and 7.5. Seawater is more basic than body fluids of the organisms living in it.

Acids and bases can be weak or strong. The weak acids, such as carbonic acid (H_2CO_3), are stingy H^+ donors. Strong acids readily give up H^+ in water. An example is the hydrochloric acid that dissociates into H^+ and Cl^- inside your stomach. The H^+ makes your gastric fluid far more acidic, which in turn activates protein-digesting enzymes.

Too much HCl can cause an *acid stomach*. Antacids taken for this condition, including milk of magnesia, release OH^- ions that combine with H^+ to reduce the pH of stomach contents.

Actually, strong acids and bases can cause severe chemical burns. That is why we are supposed to read the labels on containers of ammonia, drain cleaner, and many other common household products. That is why we are not supposed to let a car battery's sulfuric acid drip on skin.

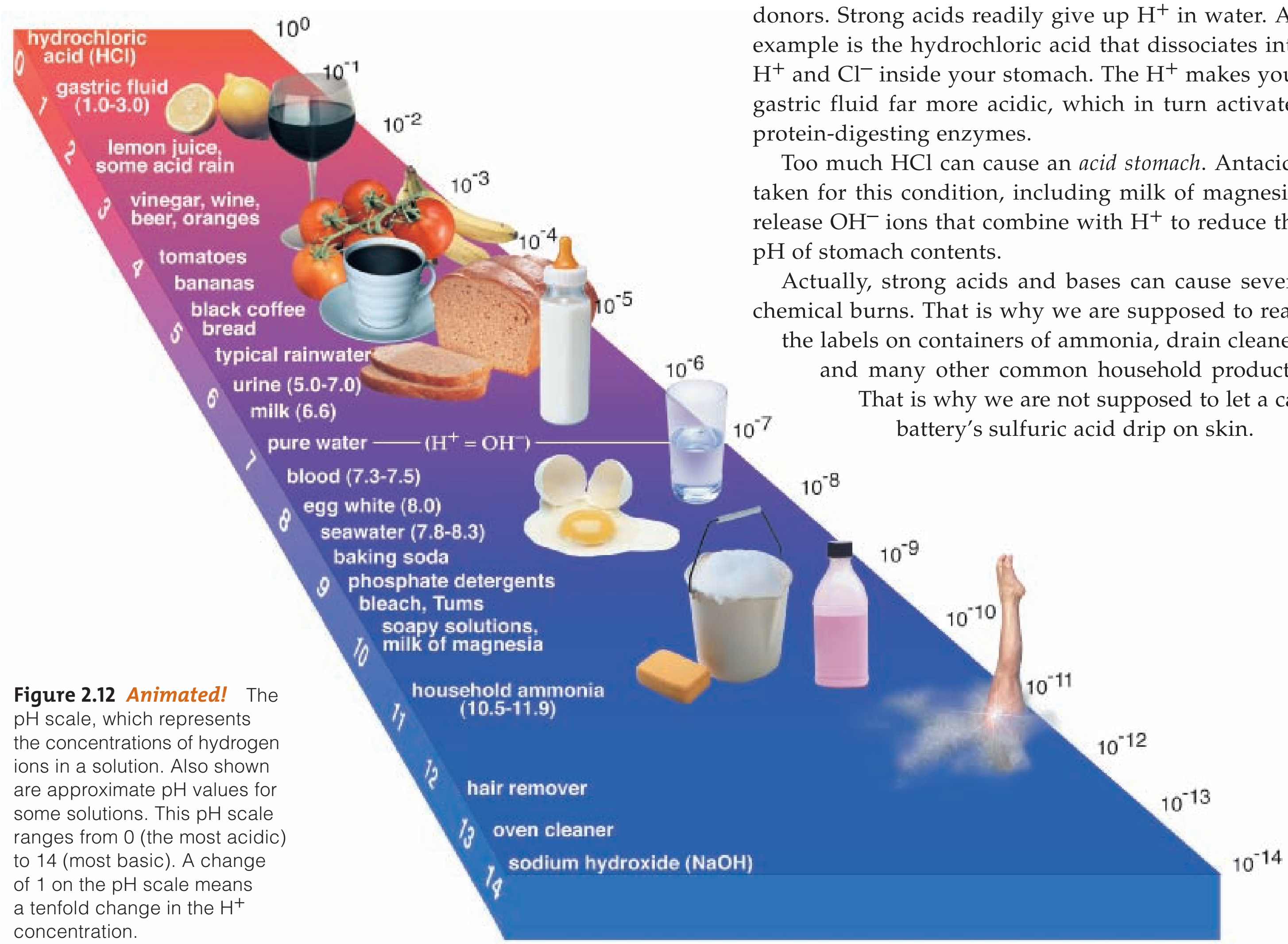


Figure 2.12 ***Animated!*** The pH scale, which represents the concentrations of hydrogen ions in a solution. Also shown are approximate pH values for some solutions. This pH scale ranges from 0 (the most acidic) to 14 (most basic). A change of 1 on the pH scale means a tenfold change in the H^+ concentration.

Figure 2.13 Emissions of sulfur dioxide from a coal-burning power plant. Airborne pollutants such as sulfur dioxide dissolve in water vapor to form acidic solutions. They are a component of acid rain.

At high concentrations, strong acids or bases that enter an ecosystem can kill organisms. For instance, fossil fuel burning and nitrogen-containing fertilizers release strong acids that lower the pH of rainwater (Figure 2.13). Some regions are sensitive to this *acid rain*. Alterations in the chemical composition of soil and water harm fishes and other organisms in these regions. We return to this topic in Section 48.2.

SALTS AND WATER

A **salt** is any compound that dissolves easily in water and releases ions *other than* H^+ and OH^-. It commonly forms when an acid interacts with a base. For example:

$$\underset{\text{HYDROCHLORIC ACID}}{HCl\ (\text{acid})} + \underset{\text{SODIUM HYDROXIDE}}{NaOH\ (\text{base})} \rightleftharpoons \underset{\text{SODIUM CHLORIDE}}{NaCl\ (\text{salt})} + H_2O$$

NaCl, the salt product of this reaction, dissociates into sodium ions (H^+) and chloride ions (Cl^-) when it is dissolved in water. Many of the ions that are released when salts dissolve in fluid are important components of cellular processes. For example, ions of sodium, potassium, and calcium are essential for nerve and muscle cell functions. They also help plant cells take up water from soil.

BUFFERS AGAINST SHIFTS IN pH

Cells must respond fast to even slight shifts in pH, because excess H^+ or OH^- can alter how biological molecules function. Responses are rapid with **buffer systems**. Think of such a system as a dynamic chemical partnership between a weak acid and its salt. These two related chemicals work in equilibrium to counter slight shifts in pH. For example, if a small amount of a strong base enters a buffered fluid, the weak acid partner can neutralize excess OH^- ions by donating some H^+ ions to the solution.

Most body fluids are buffered. Why? Enzymes, receptors, and all other essential biological molecules work most efficiently within a narrow range of pH. Deviation from the range disrupts cellular processes.

Carbon dioxide, a by-product of many reactions, becomes part of a buffer system as it combines with water to form carbonic acid and bicarbonate. When the pH of human blood rises slightly, carbonic acid can neutralize the excess OH^- by releasing hydrogen ions, which combine with OH^- to form water:

$$OH^- + \underset{\text{CARBONIC ACID}}{H_2CO_3} \longrightarrow \underset{\substack{\text{BICARBONATE} \\ \text{(salt)}}}{HCO_3^-} + \underset{\text{WATER}}{H_2O}$$

When blood becomes more acidic, bicarbonate mops up excess H^+ and thus shifts the balance of the buffer system toward the acid:

$$\underset{\text{BICARBONATE}}{HCO_3^-} + H^+ \longrightarrow \underset{\text{CARBONIC ACID}}{H_2CO_3}$$

Buffer systems can neutralize only so many excess ions. With even a slight excess above that point, the pH swings widely. When the blood pH (7.3–7.5) falls even to 7, buffering fails, and the consequences can be severe. An individual may fall into a *coma*, an often irreversible state of unconsciousness. This happens in *respiratory acidosis*. Carbon dioxide accumulates, too much carbonic acid forms, and blood pH plummets. By contrast, when the blood pH increases even to 7.8, *tetany* may occur; skeletal muscles cannot be released from contraction. *Alkalosis* is a rise in blood pH that, if not reversed by medical treatment can be lethal.

Ions dissolved in fluids on the inside and outside of cells have key roles in cell function. Acidic substances release hydrogen ions, and basic substances accept them. Salts are compounds that release ions other than H^+ and OH^-.

Acid–base interactions help maintain pH, which is the H^+ concentration in a fluid. Buffer systems help maintain the body's acid–base balance at levels suitable for life.

Summary

Introduction Chemistry can help us understand the composition and behavior of the substances that make up cells, organisms, and all components of the biosphere. Table 2.1 summarizes some key chemical terms that you will encounter throughout this book.

Section 2.1 All substances consist of one or more elements. Atoms are the smallest units that still retain the element's properties. An uncharged atom consists of one or more positively charged protons, an equal number of negatively charged electrons, and (except for hydrogen) one or more neutrons, which carry no charge. Protons and neutrons occupy an atom's core region, or nucleus, and essentially account for its mass.

Section 2.2 Atoms of an element typically differ in the number of neutrons; they are isotopes. Radioisotopes are unstable, and their nucleus spontaneously decays.

Biology Now
Learn about how radioisotopes are used in a PET scan with the animation on BiologyNow.

Section 2.3 Whether one atom will interact with others depends on the number and arrangement of its electrons. Electrons occupy orbitals (volumes of space) around the atomic nucleus. The shell model for atomic structure is a diagram with successively larger circles, or shells, that keep track of all electrons in the orbitals at a given energy level.

When an atom has one or more vacancies in orbitals, it interacts with other atoms by donating, accepting, or sharing electrons (forming chemical bonds).

Biology Now
Use the animation and interaction on BiologyNow to investigate electron distribution and the shell model.

Section 2.4 Each chemical bond is an interaction between the electron structures of atoms. The main types are called ionic, covalent, and hydrogen bonds.

When an atom loses or gains one or more electrons, it becomes an ion, with a positive or a negative charge. In an ionic bond, a positive ion and a negative ion stay together by mutual attraction of their opposite charges.

Atoms often fill vacancies in their outermost orbitals by sharing one or more pairs of electrons. Two atoms share electrons equally in a *nonpolar* covalent bond. The sharing is unequal in a *polar* covalent bond, so the bond has a slight negative charge at one end and a slight positive charge at the other. The charges balance, so the participating atoms carry no net charge, overall.

In a hydrogen bond, a covalently bound hydrogen atom weakly attracts an electronegative atom that is bound in a different molecule or a different region of the same molecule.

Biology Now
Compare the types of chemical bonds in biological molecules using the animation on BiologyNow.

Section 2.5 Polar covalent bonds join three atoms in a water molecule (two hydrogen atoms and one oxygen). The polarity of the water molecule invites extensive hydrogen bonding between molecules in bodies of water. The polarity is the basis of hydrogen bonding, which gives liquid water a notable ability to resist temperature changes, to show internal cohesion, and to easily dissolve diverse polar or ionic substances. These properties of water help make life possible.

Biology Now
Explore the structure and properties of water with the animation on BiologyNow.

Table 2.1 Summary of Important Players in the Chemical Basis of Life

Element	Fundamental substance consisting of one kind of atom
Atom	Smallest unit of an element that still retains element's properties. Occupies space, has mass, and cannot be broken apart by ordinary physical or chemical means.
Proton (p^+)	Positively charged particle of the atomic nucleus
Electron (e^-)	Negatively charged particle that can occupy a volume of space (orbital) around the nucleus
Neutron	Uncharged particle of the atomic nucleus
Isotope	One of two or more forms of an element's atoms that differ in the number of neutrons in the nucleus
Radioisotope	An unstable isotope that emits particles and energy; has an unstable combination of protons and neutrons
Tracer	Molecule that incorporates one or more atoms of a radioisotope. Used with tracking devices to identify the movement or destination of the molecule or atom in a metabolic pathway, the body, or some other system
Ion	An atom that has gained or lost an electron and carries a positive or negative charge. A proton without an electron zipping around it is a hydrogen ion (H^+)
Molecule	Unit of matter in which two or more atoms of the same element, or different ones, are bonded together
Compound	Molecule of two or more different elements in unvarying proportions (e.g., water)
Mixture	Intermingling of two or more elements or compounds in proportions that usually vary
Solute	Any molecule or ion dissolved in some solvent
Hydrophilic substance	Polar molecule or molecular region that can readily dissolve in water
Hydrophobic substance	Nonpolar molecule or molecular region that strongly resists dissolving in water
Acid	Substance that releases H^+ when dissolved in water
Base	Substance that accepts H^+ when dissolved in water
Salt	Compound that releases ions other than H^+ or OH^- when dissolved in water

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Section 2.6 The pH scale is used to measure the hydrogen ion (H^+) concentration of a solution. A typical pH range is from 0 (highest H^+ concentration; most acidic) to 14 (lowest H^+ concentration; the most basic or alkaline). At pH 7, or neutrality, H^+ and OH^- concentrations are equal.

Salts are compounds that dissolve easily in water and release ions other than H^+ and OH^-. Acids release H^+ ions in water. Bases combine with them. A buffer system is a dynamic chemical partnership between a weak acid or base and its salt. The two go back and forth donating and accepting ions to counter slight shifts in pH and thus maintain a favorable pH. Most biological processes operate within a narrow pH range.

BiologyNow

Investigate the pH of common solutions with the interaction on BiologyNow.

Self-Quiz

Answers in Appendix II

1. Is this statement false: Every type of atom consists of protons, neutrons, and electrons.

2. Electrons carry a ______ charge.
   a. positive b. negative c. zero

3. A(n) ______ is any molecule to which a radioisotope has been attached for research or diagnostic purposes.
   a. ion b. isotope c. element d. tracer

4. Atoms share electrons unequally in a(n) ______ bond.
   a. ionic c. polar covalent
   b. hydrogen d. nonpolar covalent

5. In a hydrogen bond, a covalently bound hydrogen atom weakly attracts an ______ atom in a different molecule or a different region of the same molecule.
   a. electronegative b. electropositive

6. Liquid water shows ______ .
   a. polarity d. cohesion
   b. hydrogen-bonding capacity e. b through d
   c. notable heat resistance f. all of the above

7. Hydrogen ions (H^+) are ______ .
   a. the basis of pH values d. dissolved in blood
   b. unbound protons e. both a and b
   c. targets of certain buffers f. a through d

8. When dissolved in water, a(n) ______ donates H^+, and a(n) ______ accepts H^+.

9. A(n) ______ is a dynamic chemical partnership between a weak acid and its salt.
   a. ionic bond c. buffer system
   b. solute d. solvent

10. Match the terms with their most suitable description.

____ trace element	a. atomic nucleus components
____ salt	b. two atoms sharing electrons
____ covalent bond	c. any polar molecule that readily dissolves in water
____ hydrophilic substance	d. releases ions other than H^+ and OH^- when dissolved in water
____ protons, neutrons	e. makes up less than 0.001 percent of body weight

Additional questions are available on BiologyNow™

Figure 2.14 Laboratory of a typical alchemist.

Critical Thinking

1. Some molecules consist of atoms of a single element, but others are compounds. Explain which type of molecule you would expect to be more abundant in living things.

2. *Ozone* is a chemically active form of oxygen gas. High in Earth's atmosphere, a vast layer of it absorbs about 98 percent of the sun's harmful rays. Normally, oxygen gas consists of two oxygen atoms joined in a double nonpolar covalent bond: O=O. Ozone has three covalent bonds in this arrangement: O=O—O. It is highly reactive with a variety of substances, and it gives up an oxygen atom and releases gaseous oxygen (O=O). Using what you know about chemistry, explain why you think it is so reactive.

3. Some undiluted acids are less corrosive than when diluted with a little water. In fact, lab workers are told to wipe off splashes with a towel before washing. Explain.

4. Medieval scientists and philosophers called alchemists were predecessors of modern-day chemists (Figure 2.14). Many tried to transform lead (atomic number 82) into gold (atomic number 79). Why didn't they succeed?

5. David, an inquisitive three-year-old, poked his fingers into warm water in a metal pan on the stove and didn't sense anything hot. Then he touched the pan itself and got a nasty burn. Explain why water in a metal pan heats up far more slowly than the pan itself.

6. Why can water striders (Figure 2.15) and the basilisk lizard shown in Figure 1.8 walk on water?

7. Why do you think H^+ is often written as H_3O^+?

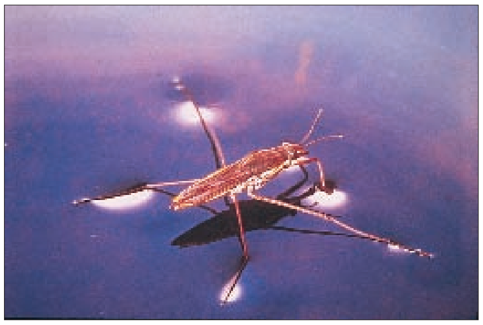

Figure 2.15 Water strider, not sinking.