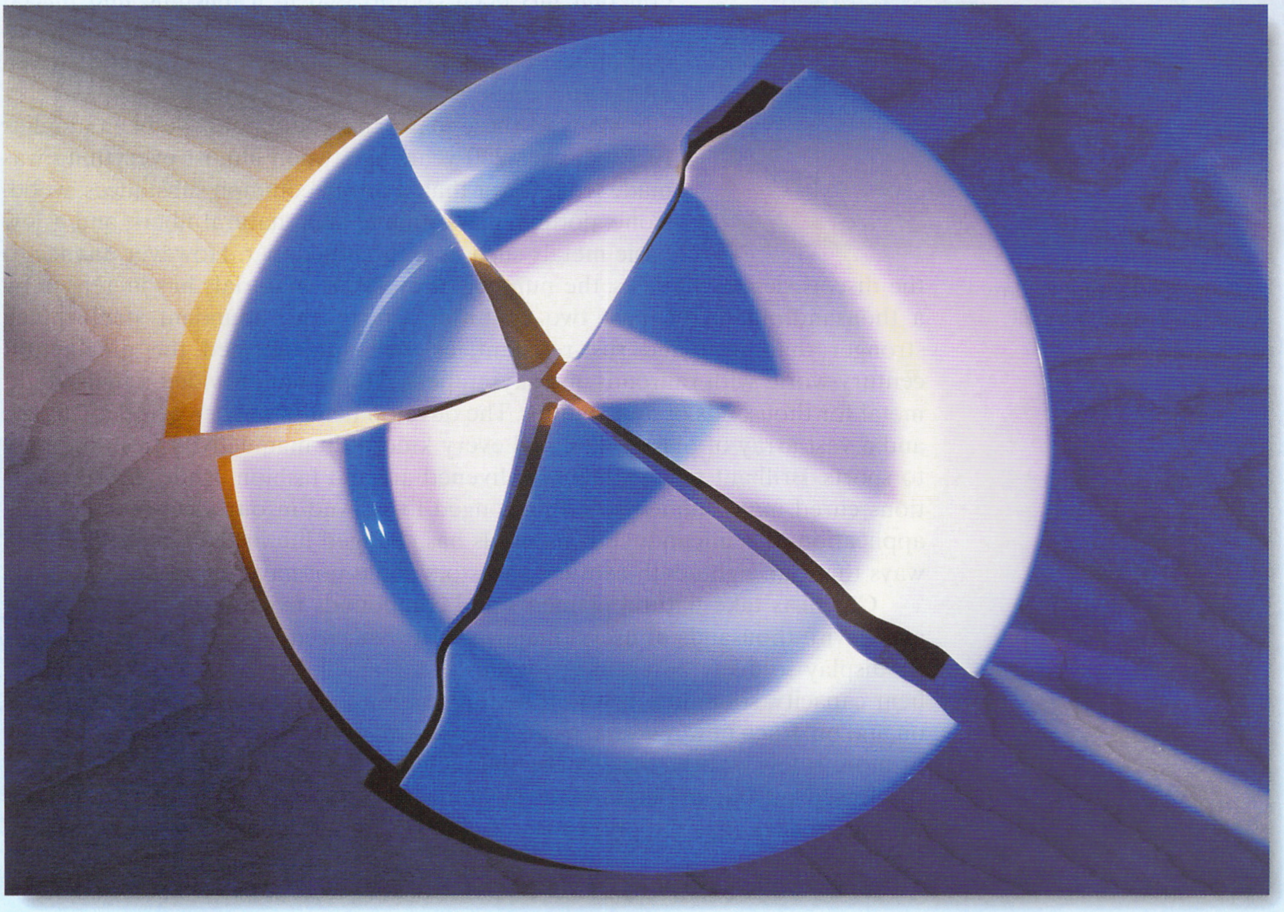
23 Chemical Bonds and Physical Properties

KEY IDEA

The properties of a material depend on the atoms from which it is made and how they are bonded together.



PHYSICS AROUND US . . . The Mystery of Dropped Objects

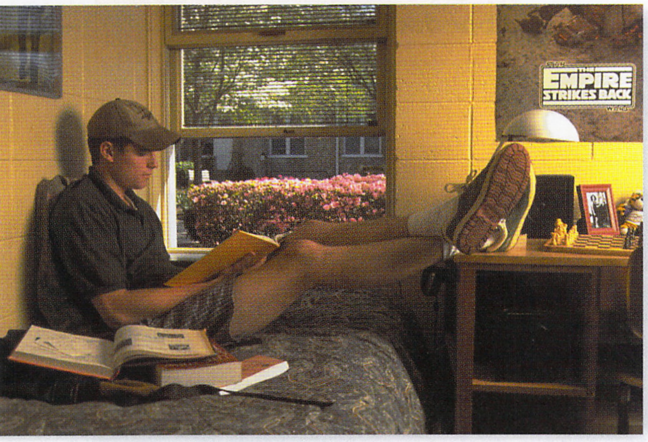
It's early morning and you're still a little groggy. As you reach for the cereal box, your hand brushes against a dish on the table, knocking it to the floor. It shatters into pieces, and your morning becomes just a little longer while you sweep them up.

Later that evening you decide to bake some cookies. As you pull one cookie tray out of the drawer, another one falls to the floor with a clatter. The metal tray doesn't break the way the dish did, however, so you just pick it up and put it back in the drawer.

Why do the plate and the tray behave so differently when they're dropped?

MATERIALS AND THE MODERN WORLD

The materials people use, perhaps more than any other facet of a culture, define the technical sophistication of their society. We speak of the most primitive human cultures as Stone Age societies and recognize Bronze Age and Iron Age peoples as progressively more advanced.



College student in a dorm room, surrounded by all sorts of materials.

Take a moment and look around your room. How many different kinds of materials do you see? The lights and windows employ glass—a brittle, transparent material. The walls may be made out of gypsum, a chalk-like mineral that has been compressed in a machine and placed between sheets of heavy paper. Your chair probably incorporates several materials, including metal, wood, woven fabric, and glues.

Many of these materials would not have been familiar to Americans 200 years ago, when almost everything was made from fewer than a dozen common substances: wood, stone, pottery, glass, animal skin, natural fibers, and a few basic metals such as iron and copper. However, thanks to the discoveries of chemists, the number of everyday materials has increased by a thousandfold in the past two centuries. Cheap and abundant steel—much stronger than iron and rust-resistant, as well—transformed the nineteenth-century world with railroads and skyscrapers. Aluminum provided a lightweight metal for thousands of applications. The development of rubber, synthetic fibers, and a vast array of plastics affected every kind of human activity from industry to sports. Brilliant new pigments enlivened art and fashion, while new medications cured many ailments and prolonged lives. And in our electronic age, the application of semiconductor materials has changed life in the United States in ways that our eighteenth-century ancestors could not have imagined.

Chemists take natural elements and compounds from earth, air, and water and devise thousands of useful materials. They succeed, in part, because materials display so many different properties: color, smell, hardness, luster, flexibility, heat capacity, solubility in water, texture, melting point, strength—the list goes on and on. Each new material holds the promise of doing some job more cheaply or more safely or otherwise better than any other.

Based on our understanding of atoms and the ways they bond together, we now realize that the properties of every material depend on three essential features:

1. The kinds of atoms of which it is made.
2. The way those atoms are arranged.
3. The way the atoms are bonded to one another.

In this chapter we look at different properties of materials and see how they relate to their atomic architecture. We examine the strength of materials—how well they resist outside forces. In the next chapter, we look at the ability of materials to conduct electricity and whether they are magnetic. Finally, in Chapter 25, we describe what are perhaps the most important new materials in modern society: the semiconductor and the microchip.

ELECTRON SHELLS AND THE CHEMICAL BOND

Think about how two atoms might interact. You know that the atom is mostly empty space, with a tiny, dense nucleus surrounded by swift electrons. If two atoms approach one another, their outer electrons—the "border guards" if you will—encounter one another first. Whatever holds two atoms together thus involves primarily those outer electrons. In fact, the outer electrons play such an important role in determining how atoms combine that they are given the special name "valence electrons." Chemical bonding often involves an exchange or sharing of valence electrons, and the number of electrons in an atom's outermost shell determines what is called the atom's "valence." Chemists often express the importance of the number of outer electrons by saying that valence represents the combining power of a given atom.

It turns out that by far the most stable arrangement of electrons—the electron configuration of lowest energy—is a completely filled outer shell. It is a fact of chemical life that different electron shells hold different numbers of electrons, which gives rise to the structure of the periodic table of the elements (see Chapter 21 and Figure 23-1). A glance at the periodic table (see Appendix C) tells us that atoms with a total of 2, 10, 18, or 36 electrons (the atoms that appear in the extreme right-hand column) have filled shells and very stable configurations. Atoms with this many electrons in their outermost shells are inert gases (also called noble gases), which do not combine readily with other materials. Indeed, helium, neon, and argon, with atomic numbers 2, 10, and 18 and thus completely filled electron shells, are the only common elements that do not ordinarily react with other elements.

Every object in nature tries to reach a state of lowest energy, and atoms are no exception. Atoms that do not have the magic number of electrons (2, 10, 18, etc.) are more likely to react with other atoms to produce a state of lower combined energy. You are familiar with this process in many other natural systems. For example, if you put a ball on top of a hill, the ball will tend to roll down to the bottom, creating a system of lower gravitational potential energy. Similarly, a compass needle tends to align itself spontaneously with Earth's magnetic field, thereby lowering its magnetic potential energy. In exactly the same way, when two or more atoms come together the electrons tend to rearrange themselves to minimize the chemical potential energy of the entire system. This situation may

1 H Hydrogen Valence, +1							2 He Helium Valence, 0
3 Li Lithium Valence, +1	4 Be Beryllium Valence, +2	5 B Boron Valence, +3	6 C Carbon Valence, +4	7 N Nitrogen Valence, −3	8 O Oxygen Valence, −2	9 F Fluorine Valence, −1	10 Ne Neon Valence, 0
11 Na Sodium Valence, +1	12 Mg Magnesium Valence, +2	13 Al Aluminum Valence, +3	14 Si Silicon Valence, +4	15 P Phosphorus Valence, −3	16 S Sulfur Valence, −2	17 Cl Chlorine Valence, −1	18 Ar Argon Valence, 0

FIGURE 23-1. The first three rows of the periodic table showing element names, atomic numbers, and principal valences. Positive valences indicate the number of electrons an element can donate or share with other elements to form compounds; negative valences indicate the number of electrons an element can accept or share to form compounds.

require that they exchange or share electrons. As often as not, that process involves rearrangements with a total of 2, 10, 18, or 36 electrons.

Chemical bonds result from any redistribution of electrons that leads to a more stable configuration between two or more atoms, especially that of a filled electron shell.

Most atoms adopt one of three simple strategies to achieve a filled shell: they give away electrons, accept electrons, or share electrons.

If the bond formation takes place spontaneously, without outside intervention, energy is released in the reaction. The burning of wood or paper (once their temperature has been raised high enough) is a good example of this sort of process. The heat you feel when you put your hands toward a fire derives ultimately from the chemical potential energy that is given off as electrons and atoms are reshuffled. Alternatively, atoms may be pushed into new configurations by adding energy to systems. Much of industrial chemistry, from the smelting of iron to the synthesis of plastics, operates on this principle.

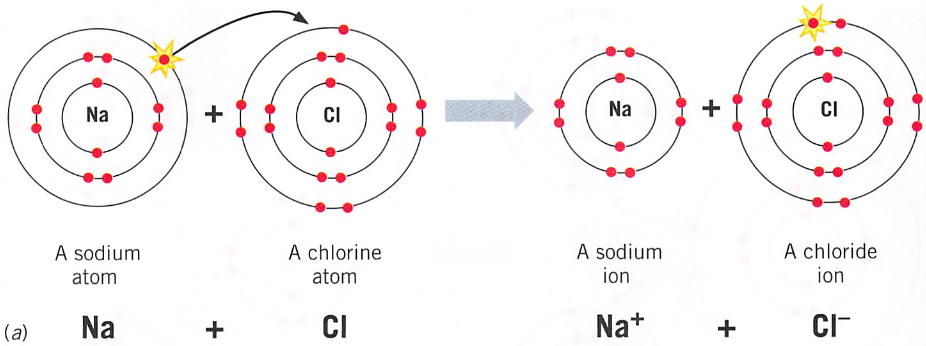
Types of Chemical Bonds

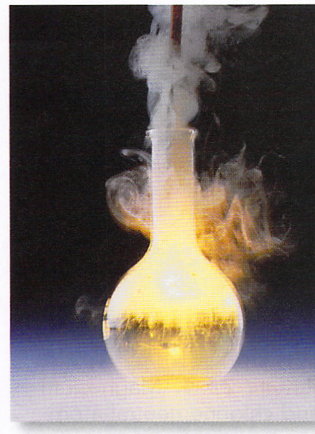
Atoms link together by three principal kinds of chemical bonds—ionic, metallic, and covalent—all of which involve redistributing electrons between atoms. In addition, three types of attractive forces—polarization, van der Waals interactions, and hydrogen bonding—can result from the shifting of electrons within their atoms or groups of atoms. Each type of bonding or attraction corresponds to a different way of rearranging electrons, and each produces distinctive properties in the materials it forms.

Ionic Bonds We have seen that atoms with 2, 10, 18, or 36 electrons are particularly stable. By the same token, atoms that differ from these magic numbers by only one electron in their outer shells are particularly reactive—in effect, they are "anxious" to fill or empty their outer shells. Such atoms tend to form **ionic bonds,** chemical bonds in which the electric force between two oppositely charged ions holds the atoms together.

Ionic bonds often form as one atom gives up an electron while another receives it. For example, sodium (a soft, silvery white metal) has 11 electrons in an electrically neutral atom—2 in the lowest shell, 8 in the next, and a single electron in its outer shell. Sodium's best bonding strategy, therefore, is to lose one electron. Element 17, chlorine (a yellow-green toxic gas), is one electron shy of a filled shell. Highly corrosive chlorine gas reacts with almost anything that can give it an extra electron. When you place sodium in contact with chlorine gas, the result is predictable: in a fiery reaction, each sodium atom donates its extra electron to a chlorine atom (Figure 23-2).

In the process of this vigorous electron exchange, atoms of sodium and chlorine become electrically charged—they become ions. Neutral sodium has 11 positive protons in its nucleus, balanced by 11 negative electrons. By losing an electron, sodium becomes an ion with one unit of positive charge, shown as Na^+ in Figure 23-2. Similarly, neutral chlorine has 17 protons and 17 electrons. The addition of an extra negative electron creates a chloride ion with one unit of negative charge, shown as Cl^- in the figure. The mutual electrical attraction of positive sodium and negative chloride ions is what forms the ionic bonds between





(b)

Figure 23-2. (*a*) Sodium, a highly reactive element, readily transfers its single valence electron to chlorine, which is one electron shy of the "magic" number 18. In these diagrams, electrons are represented as dots in shells around a nucleus. (*b*) The result of this fiery reaction is the ionic compound sodium chloride, or ordinary table salt.

sodium and chlorine. The resulting compound, sodium chloride (common table salt), has properties totally different from either sodium or chlorine.

Under normal circumstances, sodium and chloride ions lock together into a crystal, a regular arrangement of atoms such as the one shown in Figure 23-3. Alternating sodium and chloride ions form an elegant repeating structure in which each Na^+ is surrounded by six Cl^-, and vice versa.

Ionic bonds may involve more than a single electron transfer. For example, element 12, magnesium, donates two electrons to oxygen, which has six valence electrons. In the resulting compound, MgO (magnesium oxide), both atoms have stable filled shells of eight electrons, and the ions, Mg^{2+} and O^{2-}, form a strong ionic bond. Ionic bonds involving the negative oxygen ion O^{2-} and positive ions, such as aluminum (Al^{3+}), magnesium (Mg^{2+}), and iron (Fe^{2+} or Fe^{3+}), are found in many everyday objects: in most rocks and minerals, in china and glass, and in bones and egg shells.

The ionic bonds in these compounds can be very strong, but only in certain ways. You can picture how this works by thinking about Tinkertoys. A Tinkertoy structure can be quite strong: when assembled, it is difficult to break one apart by just pushing in the directions of the sticks. But Tinkertoy bonds break easily if you twist or snap the sticks. In the same way, ionic bonds hold atoms together, but if for some reason the atoms become displaced, the bond can't hold them very well. As a consequence, ionic-bonded materials such as rock, glass, and eggshells are usually quite brittle. These materials are strong in the sense that you can pile a lot of weight on them. But once they shatter and the ionic bonds are broken, they can't be put back together again. This is why the dish described in Physics Around Us (page 491) shattered when it hit the floor.

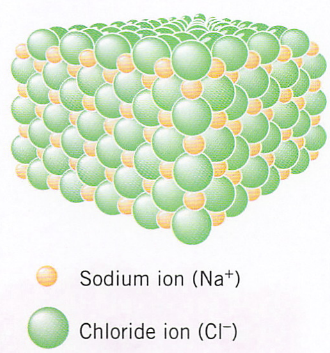


Figure 23-3. The atomic structure of a sodium chloride crystal consists of a regular pattern of alternating sodium and chloride ions.

Ionic Bonding of Three Atoms

EXAMPLE 23-1

Magnesium chloride, which plays an important role in some types of batteries, is an ionic-bonded compound with one part magnesium to two parts chlorine ($MgCl_2$). How are the electrons arranged in this compound?

REASONING: From the periodic table (see Appendix C), magnesium and chlorine are elements 12 and 17, respectively. Magnesium, therefore, has 10 electrons (2 + 8) in inner shells and 2 valence electrons. Chlorine has 10 electrons (2 + 8)

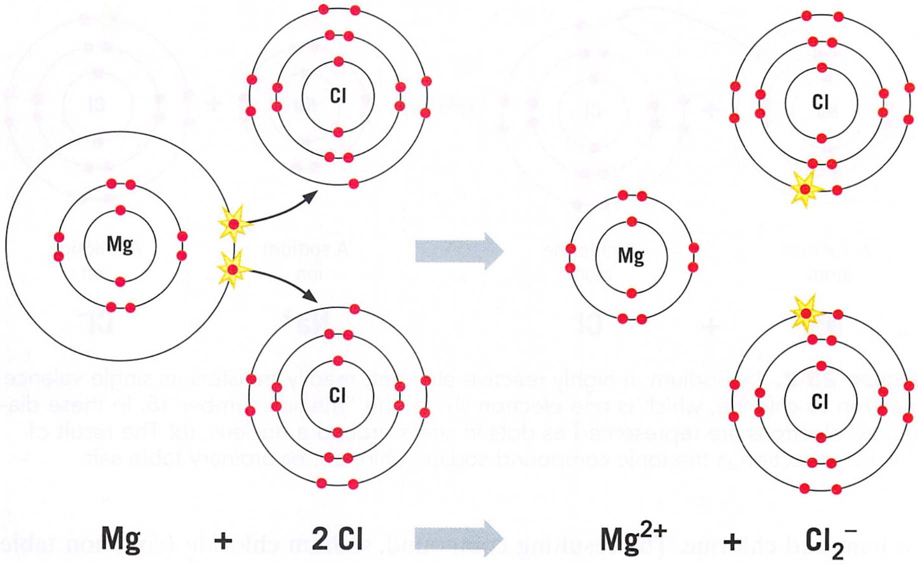


FIGURE 23-4. Magnesium and chlorine neutral-atom electron configurations (*left*), and their configurations after electrons have been transferred from the magnesium to the chlorine atoms (*right*).

in its inner shells and 7 electrons in the outer one, meaning that it is 1 electron short of a filled outer shell (Figure 23-4).

SOLUTION: Magnesium has two electrons to give and chlorine seeks one electron to achieve stable filled outer shells. Thus magnesium gives one electron to each of two chlorine atoms, and the resulting Mg^{2+} ion attracts two Cl^- ions to form $MgCl_2$.

Metallic Bonds Atoms in an ionic bond transfer electrons directly—electrons are on permanent loan from one atom to another. Atoms in a metal also give up electrons, but they use a very different bonding strategy. In a **metallic bond,** electrons are redistributed so that many atoms share them.

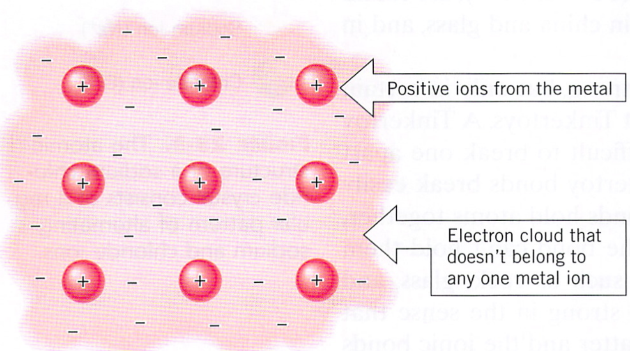


FIGURE 23-5. Metallic bonding, in which a bond is created by the sharing of electrons among several metal atoms.

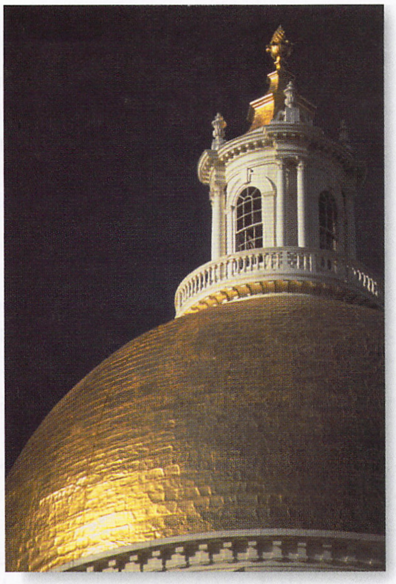
Sodium metal, for example, is made up entirely of individual sodium atoms. All of these atoms begin with 11 electrons, but they release one to achieve the more stable 10-electron configuration. The extra electrons move away from their parent atoms to float around the metal, forming a kind of sea of negative charge. In this negative electron sea, the positive sodium ions adopt a regular crystal structure, as shown in Figure 23-5.

You can think of the metallic bond as one in which each atom shares its outer electron with all the other atoms in the system. Picture the free electrons as a kind of loose glue in which the metal ions are placed. In fact, the idea of a metal as a collection of marbles (the ions) in a sea of stiff, gluelike liquid provides a useful analogy.

Metals are formed by almost any element or combination of elements in which large numbers of atoms share electrons to achieve a more stable electron arrangement. Metals are characterized by their shiny luster and ability to conduct electricity. These properties are due to the ability of the loosely held electrons to interact with electromagnetic waves and fields. Some metals, such as aluminum, iron, and copper, are familiar from everyday experience. However, many elements can form a metallic state when the conditions are right, including some that we normally think of as gases, such as hydrogen or oxygen at very high pressure. (The planet Jupiter consists mostly of metallic hydrogen.) In fact, the great majority of chemical elements can occur in a metallic state. In addition, two or more elements can combine to form a metal "alloy," such as brass (a mixture of copper and zinc) or bronze (an alloy of copper and tin). Modern specialty-steel alloys often contain more than half a dozen different elements in carefully controlled proportions.

The special nature of the metallic bond explains many of the distinctive properties we observe in metals. If you attempt to deform a metal by pushing on the marble-and-glue bonding system, atoms will gradually rearrange themselves and come to some new configuration—the metal is malleable. It's hard to break a metallic bond just by pushing or twisting because the atoms are able to rearrange themselves. Thus, when you hammer on a piece of metal you leave indentations but do not break it, in sharp contrast to what happens when you hammer on a ceramic plate.

In the next chapter we examine more closely the electric properties of materials held together by metallic bonds. We see that this particular kind of bond produces materials through which electrons—electric current—can flow.



Gold metal is so soft and malleable that it can be hammered paper-thin and applied to surfaces—an art known as gilding. The dome of the State House in Boston, Massachusetts, is gilded in this manner.

Covalent Bonds In an ionic bond, one atom donates electrons to another on more or less permanent loan. In a metallic bond, on the other hand, atoms share some electrons throughout the material. In between these two types of bond is the extremely important **covalent bond,** in which well-defined clusters of neighboring atoms, called *molecules,* share electrons. These strongly bonded groups may consist of anywhere from two atoms to many millions.

The simplest covalently bonded molecules contain two atoms of the same element, such as the diatomic gases hydrogen (H_2), nitrogen (N_2), and oxygen (O_2). In the case of hydrogen, for example, each atom has a relatively unstable single electron. Two hydrogen atoms can pool their electrons, however, to create a more stable two-electron arrangement. The two hydrogen atoms must remain close to one another for this sharing to continue, so a chemical bond is formed, as shown in Figure 23-6. Similarly, two oxygen atoms, each with eight electrons (six in the outer shell), share two pairs of electrons.

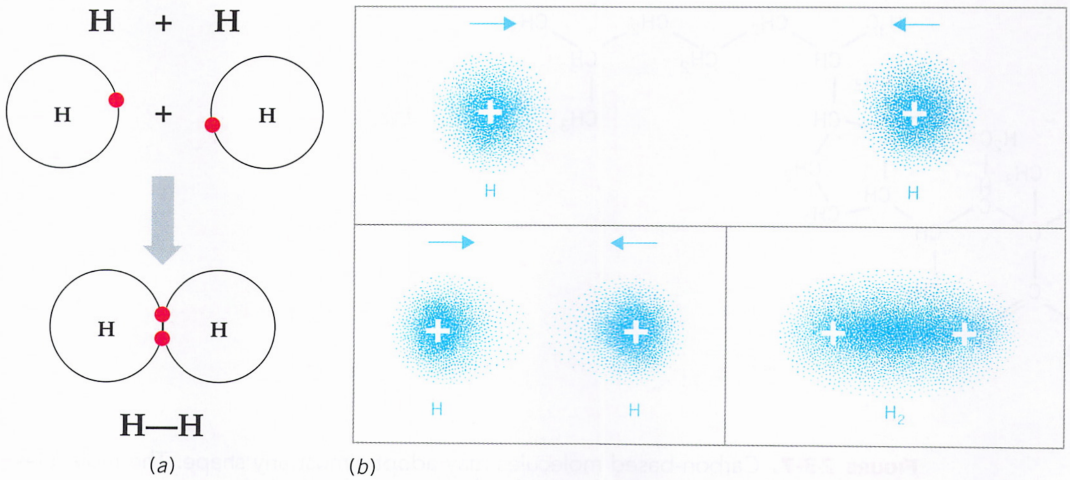
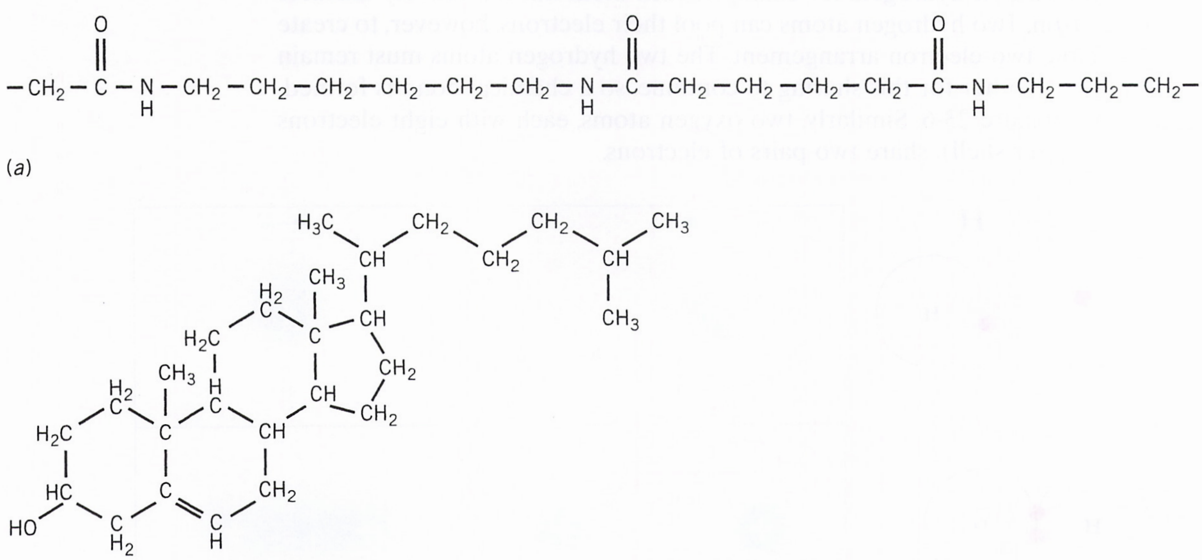


FIGURE 23-6. Two hydrogen atoms become an H_2 molecule by sharing each of their electrons in a covalent bond. This bonding may be represented schematically in a dot diagram (*a*) or by the merging of two atoms with their electron clouds (*b*).

Hydrogen, oxygen, nitrogen, and other covalently bonded molecules have lower chemical potential energy than isolated atoms have because electrons are shared. The negative electrons are attracted to two positive nuclei, not one, which reduces their potential energy. These molecules are less likely to react chemically than are the isolated atoms.

The most fascinating of all elements that form covalent bonds is carbon, which forms the backbone of all life's essential molecules. Carbon, with two electrons in its inner shell and four in its outer shell, presents a classic case of a half-filled shell. When carbon atoms approach one another, therefore, a real question arises as to whether they ought to accept or donate four electrons to achieve a more stable arrangement. You could imagine, for example, a situation where some carbon atoms give four electrons to their neighbors, while other carbon atoms accept four electrons, to create a compound with strong ionic bonds between C^{4+} and C^{4-}. Alternatively, carbon might become a metal in which every atom releases four electrons into an extremely dense electron sea. But neither of these things happens.

In fact, the strategy that lowers the energy of the carbon–carbon system the most is for the carbon atoms to share their outer electrons. Once bonds between carbon atoms have formed, the atoms have to stay close to each other for the sharing to continue. Thus the bonds generated are just like the bond in the hydrogen molecule. The case of carbon is unusual, however, because the shape of its electron shells allows a single carbon atom to form covalent bonds with up to four other atoms by sharing one of its four valence electrons with each. A *single bond* (shown as C—C) forms when one electron from each atom is shared, while a *double bond* (shown as C=C) results when two electrons from each atom are shared.



(b)

FIGURE 23-7. Carbon-based molecules may adopt almost any shape. The molecules may consist of long, straight chains of carbon atoms that form fibrous materials such as nylon (*a*), or they may incorporate complex rings and branching arrangements that form lumpy molecules such as cholesterol (*b*).

By forming bonds among several adjacent carbon atoms, you can make rings, long chains, branching structures, planes, and three-dimensional frameworks of carbon in almost any imaginable shape. There is virtually no limit to the complexity of molecules you can build from such carbon–carbon bonding (Figure 23-7). So important is the study of carbon-based molecules that chemists have given it a special name, *organic chemistry.*

Connection

Chemical Bonds in the Human Body

All the molecules in your body and in every other living thing are held together at least in part by covalent bonds in carbon chains. Covalent bonds also drive much of the chemistry in the cells of your body and play a role in holding together the DNA molecules that carry your genetic code. It would not be too much of an exaggeration to say that the covalent bond is the bond of life.

Let's consider one simple example. The food you eat is converted in your body to molecules of a sugar called glucose, with the chemical formula $C_6H_{12}O_6$ (that is, it contains 6 carbon atoms, 6 oxygen atoms, and 12 hydrogen atoms). When cells in your body need energy—for sending nerve impulses, for activating muscles, or for any bodily activity—these glucose molecules are burned as fuel in a combustion reaction known as respiration. In the end, respiration is very similar to the combustion of wood or gasoline, producing the products carbon dioxide and water. This reaction releases energy stored in the covalent bonds of the glucose molecule. You breathe air in order to obtain oxygen for respiration and to get rid of the water vapor and carbon dioxide produced. ●

Develop Your Intuition: The Element of Life

Life on Earth is based on the properties of the element carbon. Looking at the first three rows of the periodic table in Figure 23-1, do you see any other candidate elements that might form the basis of life elsewhere?

In the periodic table, the place to look for similar elements is in the same vertical column. Right underneath carbon in the periodic table is the element silicon, with the same arrangement of four electrons in its outer shell. Silicon forms a wide variety of compounds, as carbon does. Many of these silicon compounds are directly analogous to carbon compounds.

The major difference between carbon and silicon is that silicon has three shells of electrons instead of carbon's two shells. The result is that the electrons in silicon are spread out more in space and thus form longer and weaker bonds than carbon does. These weaker bonds make silicon compounds more reactive than similar carbon compounds. For example, low-weight carbon molecules are used as fuel in gasoline while higher-weight carbon molecules form waxes. However, the analogous low-weight silicon compounds are very dangerous: they can ignite or explode spontaneously in air. In addition, silicon compounds do not have high molecular weights: the silicon bonds are too weak for more than about eight atoms in a chain. Silicon forms a wide variety of compounds with oxygen and is present in most rocks and minerals; the oxygen atoms help form stable bonds between the silicon atoms and atoms of other elements.

Even with these differences, science fiction writers sometimes imagine life forms based on silicon, due to the great variety of compounds it can form. The astronomer and science fiction writer Fred Hoyle wrote a novel called *The Black Cloud,* in which he imagined an intelligent being in the form of a huge interstellar cloud of gas consisting of carbon and silicon compounds. But most biologists with interests in other life forms agree that silicon as the basis for life is likely to remain fiction rather than fact.

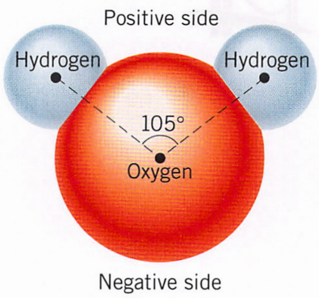


FIGURE 23-8. The water molecule and its polarity.

Polarization and Hydrogen Bonds Ionic, metallic, and covalent bonds form strong links between individual atoms within molecules. However, molecules also experience forces that hold one molecule to another. In many cases, the electric forces are such that, although the molecule by itself is electrically neutral, one part of the molecule has more positive or negative charge than another. For example, in water the electrons tend to spend more time around the oxygen atom than around the hydrogen atoms. This uneven electron distribution has the effect of making the oxygen side of the water molecule more negatively charged and the two Mickey Mouse ears of the hydrogen atoms more positively charged (Figure 23-8). Atomic clusters of this type, with a positive and negative end, are called *polar molecules* (see Chapter 16).

The electrons of an atom or molecule brought near a polar molecule such as water tend to be pushed away from the negative side and shifted toward the positive side. Consequently, the side of an atom facing the negative end of a polar molecule becomes slightly positive. This subtle electron shift, called *polarization,* gives rise to an electrical attraction between the negative end of the polar molecule and the positive side of the neighboring molecule. This electron movement thus creates an attraction between molecules, even though the atoms and molecules in this scheme all may be electrically neutral. One of the most important consequences of forces due to polarization is the ability of water to dissolve many materials. Water, made up of strongly polar H_2O molecules, exerts forces that make it easier for ions such as Na^+ and Cl^- to separate from one another and surround themselves with water molecules instead. The result is that salt dissolves in water.

A process related to the forces of polarization leads to the *hydrogen bond.* Hydrogen bonds are weak bonds that may form after a hydrogen atom links to an atom of certain other elements (oxygen or nitrogen, for example) by a covalent bond. Because of the kind of rearrangement of electric charge just described, hydrogen may become polarized and develop a slight positive charge that attracts another atom to it. You can think of the hydrogen atom as a kind of bridge in this situation, causing a redistribution of electrons that in turn holds the larger atoms or molecules together. Individual hydrogen bonds are weak, but in many molecules they occur repeatedly and therefore play a major role in determining the molecule's shape and function. Note that while all hydrogen bonds require hydrogen atoms, not all hydrogen atoms are involved in hydrogen bonds.

Hydrogen bonds are common in virtually every biological substance, from everyday materials such as wood, plastics, silk, and candle wax to the complex structures of every cell in your body. Hydrogen bonds in every living thing link the two sides of the DNA double helix together, although the sides themselves are held together by covalent bonds. Ordinary egg white is made from molecules whose shape is determined by hydrogen bonds; when you heat the material—when you fry an egg, for example—you break these hydrogen bonds. As a result, the molecules rearrange themselves so that instead of a clear gelatinous liquid you have a white solid.

Van der Waals Forces Hydrogen bonds exist because atoms or molecules can become polarized as their electrons shift to one side or another and thus create local electric charges. In the molecules we've discussed so far, that electric charge is more or less permanently locked into polar molecules in a fixed or static arrangement. Another force between molecules, called the *van der Waals force,* results from the polarization of electrically neutral atoms or molecules that are not themselves polar.

When two atoms or molecules are brought near one another, every part of one atom or molecule feels an electric force exerted on it by all parts of the other. For example, an electron in one atom is repelled by the electrons of an adjacent atom, but is attracted to the adjacent nucleus. The net result of these forces exerted on the electron may be a temporary shift of the electron. The same thing happens to every electron in any nearby atom or molecule, and the net result is that every electron is constantly shifting because of the presence of others.

What is remarkable is that sometimes this mutual, dynamic deformation can give rise to a net attractive force. In compounds where this happens, even if all the molecules are neutral and nonpolar, the sum of attractive forces wins out over repulsive forces and weak bonds are formed. This weak force that binds two atoms or molecules together is the van der Waals force.

If you take a piece of clay and rub it between your fingers, your fingers pick up a slick coating of the material, even though the clay crumbles easily. The reason for this behavior is that the clay is made up of sheets of atoms. Within each sheet, atoms are held together by strong ionic and covalent bonds. However, one sheet is held to another by comparatively weak van der Waals forces. This situation is not unlike the way a stack of photocopying paper will stick together on a dry day. Each sheet of paper is strong, but the stack of paper sticks together because of much weaker electrostatic forces. It is easy to pull the stack apart, but very difficult to rip the stack in two. When you crumble clay in your fingers, therefore, you are breaking weak van der Waals forces between layers but preserving the stronger bonds that hold each layer together. The clay stains on your hands are thin sheets of atoms, held together by ionic and covalent bonds; the crumbling is due to the breaking of the van der Waals bonds.

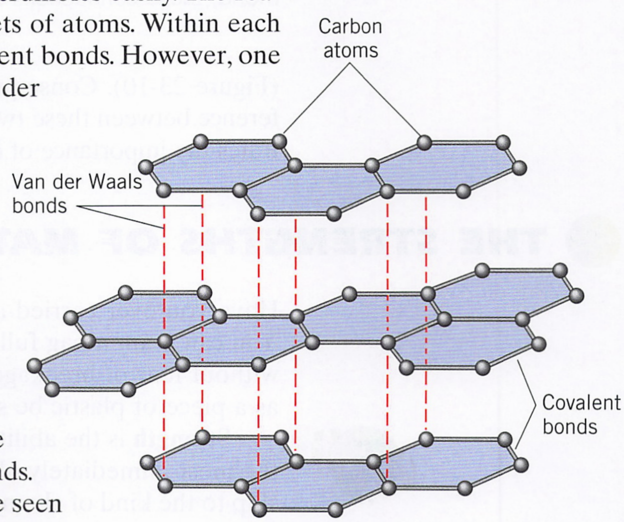


FIGURE 23-9. Graphite, a form of carbon that serves as the lead in your pencil, contains layers of carbon atoms strongly linked to one another by covalent bonds (represented by double solid lines). The separate layers are held together by much weaker van der Waals bonds (represented by dashed lines).

Many other examples of van der Waals forces can be seen in everyday life. If you rub talcum powder on your body, for example, you use a layered material not unlike the clay just discussed. Similarly, when you write with a lead (really graphite, a form of carbon) pencil, van der Waals–bonded layers of graphite (Figure 23-9) are transferred from the pencil to the paper. As you draw the pencil across the paper, you break the van der Waals forces and leave behind dark graphite sheets on the paper. Van der Waals forces also link molecules in many everyday liquids and soft solids, from candle wax to Vaseline and other petroleum products.

Contrast the behavior of the graphite in your pencil to the behavior of a diamond, which is also made from pure carbon. In a diamond, all the carbon atoms are locked together by covalent bonds in a complex three-dimensional array

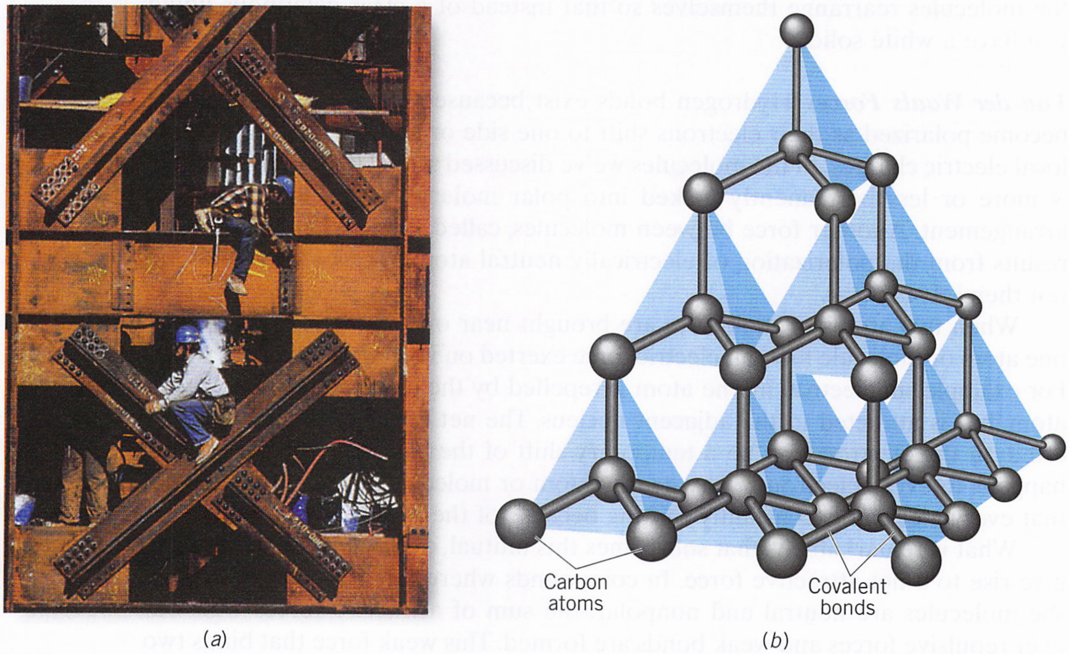


FIGURE 23-10. The girder framework of a skyscraper (*a*) and the crystal structure of diamond (*b*) are both strong because of numerous very strong connections. In diamond, these connections are covalent carbon–carbon bonds.

(Figure 23-10). Consequently, diamond is the hardest material known. The difference between these two materials, both made from exactly the same atoms, illustrates the importance of chemical bonds in determining how a substance behaves.

THE STRENGTHS OF MATERIALS

Have you ever carried a heavy load of groceries in a thin plastic grocery bag? You can cram a bag full of heavy bottles and cans and lift it by its thin handles without fear of breakage. How can something as light, flexible, and inexpensive as a piece of plastic be so strong?

Strength is the ability of a solid to resist changes in shape. Strength is one of the most immediately obvious material properties and it bears a direct relationship to the kind of chemical bonds present. A strong material must be made with strong chemical bonds. By the same token, a weak material, like a defective chain, must have weak links between some of its atoms. While no type of bond or attraction is universally stronger than the other kinds, van der Waals forces are generally the weakest. Any material with van der Waals forces will have one or more particularly soft directions of bonding and you will probably be able to pull the material apart with your hands. You experience this softness whenever you use baby powder, graphite lubricant, or soap.

By contrast, many strong materials, such as rocks, glass, and ceramics, are held together primarily by ionic bonds. The next time you see a building under construction, look at the way beams and girders link diagonally to form a rigid framework. Chemical bonds in strong materials do the same thing. A three-dimensional network of ionic bonds in these materials holds them together like a framework of steel girders.

The strongest materials we know, however, incorporate long chains and clusters of carbon atoms held together by covalent bonds. The extraordinary strength of natural spider webs, synthetic Kevlar (used to make bulletproof vests), diamonds, and your plastic shopping bag all stem from the strength of covalent bonds to carbon atoms.

Kevlar is lightweight, flexible, and strong enough to stop a bullet; it is the material from which bulletproof vests are made.

Different Kinds of Strength

Every material is held together by the bonds between its atoms. When an outside force is applied to a material, the atoms must shift their positions in response. The bonds stretch and compress, and an equal and opposite force is generated inside the material to oppose the force that is imposed from the outside, in accordance with Newton's third law of motion. The strength of a material is thus related to the size of the force it can withstand when it is pushed or pulled.

Material strength is not a single property because there are different ways of placing an object under stress. Scientists and engineers recognize three very different kinds of strength when characterizing a material:

1. Its ability to withstand crushing (*compressive strength*).
2. Its ability to withstand pulling apart (*tensile strength*).
3. Its ability to withstand twisting (*shear strength*).

Your everyday experience should convince you that these three properties are often quite independent. For example, a loose stack of bricks can withstand crushing pressures—you can pile tons of weight on it without having the stack collapse because each brick pushes down on the one beneath it but cannot exert a strong force in the sideways direction. But the same stack of bricks has little resistance against twisting; indeed, a child can topple it. A rope, on the other hand, is extremely strong when pulled, but has little strength under twisting or crushing.

The point at which a material is no longer strong enough to resist external forces and begins to bend, break, or tear is called its *elastic limit.* We see examples of this behavior every day. When you break an egg, crush an aluminum can, overstretch a rubber band, or fold a piece of paper, you exceed an elastic limit and permanently change the object. When the materials in your body exceed their elastic limit, the consequences can be catastrophic. Our bones if put under too much stress may break, while our arteries if put under pressure that is too high may rupture in an aneurysm.

A material's strength is a result of the type and arrangements of chemical bonds. Think about how you might design a structure using Tinkertoys that would be strong under crushing, pulling apart, or twisting. The strongest arrangement would have lots of short sticks with triangular patterns. Nature's strongest structure, diamond, adopts this strategy: it is exceptionally strong under all three kinds of stress because of its three-dimensional framework of strong carbon–carbon bonds (see Figure 23-10). Glass, ceramics, and most rocks, which also feature rigid frameworks of chemical bonds, are relatively strong. However, many plastics, such as the one used in your shopping bag, have strong bonds in only one direction and thus are strong when stretched, but have little strength when twisted or crushed. Materials with layered atomic structures, in which planes of atoms are arranged like a stack of paper, are generally strong when squeezed, but quite weak under other stresses. Thus the strength of a material depends on the kinds of atoms in it, the way they are arranged, and the kinds of chemical bonds that hold the atoms together.

Develop Your Intuition: Modern Building Materials

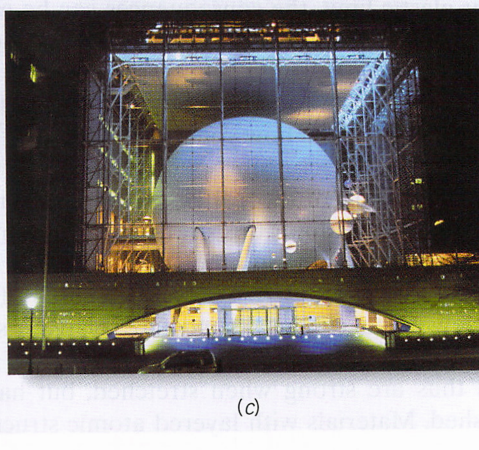
Architects specify a wide variety of different materials in contemporary building design. If you were planning a modern museum or concert hall that needed to be elegant as well as strong, would you be more likely to choose concrete, steel, or glass as building material?

Actually, all three materials are widely used in modern architecture, which is testimony to the progress of materials science over the last few decades. Steel has long been the workhorse of modern building design, but when used with imagination, steel shows elegance as well as strength. The John Hancock Center in Chicago, completed in 1970, shows one aspect of steel as an elegant material; another is the Pompidou Center, in Paris, finished in 1977. In the Hancock Building, many of the steel supports are in plain view on the building's exterior, instead of being hidden inside.

Concrete is usually considered to be a material used for its brute strength in withstanding compression, as in most large dams around the world. However, concrete can also soar in seemingly delicate arches, such as the Gateway

(*a*)

(*b*)



(*c*)

Building materials: (*a*) steel in the John Hancock Building, Chicago; (*b*) concrete in the Sydney Opera House, Australia; (*c*) glass in the Rose Center of the American Museum of Natural History, New York.

Arch in St. Louis, built in 1965 from steel reinforced with concrete. Perhaps even more graceful is the roof of the Opera House in Sydney, Australia, completed in 1973, which consists of concrete arches joined together to form a double series of overlapping shells.

Glass has long been used for windows, but over the years it has been strengthened for use as walls in large skyscrapers. Recent advances have led to glass that retains its transparency as well as gaining strength, as in the Rose Center for Earth and Space in New York City, which opened in 2000.

Modern materials science has advanced to the point that strength and elegance can be combined in steel, concrete, or glass. The only limit is the vision and imagination of the architect.

Composite Materials

As we have seen, some materials resist compression better than tension, while other materials can withstand great tensile forces but crack easily under compression. *Composite materials* combine the properties of two or more materials. The strength of one of the constituents is used to offset the weakness of another, resulting in a material whose strengths are greater than those of any of its components. Plywood, one of the most common composite materials, consists of thin wood layers glued together with alternating grain direction. The weakness of a single thin sheet of wood is compensated by the strength of the neighboring sheets. Not only is plywood much stronger than a solid board of the same dimension, but it can also be produced from much smaller trees by slicing thin layers of wood off a rotating log, like removing paper from a roll.

Reinforced concrete is used in construction for its combined strength in resisting both tension and compression.

Reinforced concrete is a common composite material in which steel rods (with great tensile strength) are embedded in a concrete mass (with great compressive strength). A similar strategy is used in fiberglass, formed from a cemented mat of glass fibers. New carbon-fiber composites provide extraordinarily strong and lightweight materials for industry and sports applications.

The modern automobile features a wide variety of composite materials. Windshields of safety glass are layered to resist shattering and reduce sharp edges in a collision. Tires are intricately formed from rubber and steel belts for strength and durability. Car upholstery commonly mingles natural and artificial fibers, and dashboards often employ complex laminated surfaces. The bodies of many cars are formed from fiberglass or other molded lightweight composite. And, as we see later, all of a modern automobile's electronics, from radio to ignition, depend on semiconductor composites of extraordinary complexity.

Physics in the Making

The Discovery of Nylon

Nature's success in making strong, flexible fibers inspired scientists to try the same thing. American chemist Wallace Carothers (1896–1937) began thinking about polymer formation while a graduate student in the 1920s. At the time, no one was sure how natural fibers formed, or what kinds of chemical bonds were involved. Carothers wanted to find out.

The chemical company DuPont took a gamble by naming Carothers head of its new fundamental research group in 1928. No pressure was placed on him to produce commercial results, but within a few years his team had developed the synthetic rubber neoprene. By the mid-1930s they had devised a variety of extraordinary polymers (see Chapter 9), including nylon, the first human-made fiber. Carothers also demonstrated conclusively that polymers in nylon are covalently bonded chains produced from small molecules, each of which has six carbon atoms, linked together.

DuPont made a fortune from nylon and related synthetic fibers. Nylon is inexpensive to manufacture and has many advantages over natural fibers. It can be melted and squeezed out of spinnerets to form strands of almost any desired size—for example, threads, rope, surgical sutures, tennis racket strings, and paintbrush bristles. These fibers can be made smooth and straight, like fishing line, or rough and wrinkled, like wool, to vary the texture of fabrics. Nylon fibers can also be kinked with heat to provide permanent folds and pleats in clothing. The melted polymer can even be injected into molds to form durable parts such as tubing or zipper teeth.

Sadly, Wallace Carothers did not live to see the impact of his extraordinary discoveries. Suffering from increasingly severe bouts of depression and convinced that he was a failure as a scientist, Carothers took his own life in 1937, just a year before the commercial introduction of nylon.

CHEMICAL REACTIONS AND THE FORMATION OF CHEMICAL BONDS

Atoms and small molecules can come together to form larger molecules. Large molecules can break up into atoms and/or smaller molecules. We call these processes **chemical reactions.** When we take a bite of food, light a match, wash our hands, or drive a car, we initiate chemical reactions. Earth's chemical reactions include rock formation and rock weathering, atmospheric weathering, soil formation, water erosion, and the cycling of elements around the world. Every moment of every day, countless chemical reactions in every cell of our bodies sustain life.

All chemical reactions involve the rearrangement of the atoms in elements and compounds, as well as the rearrangement of electrons to break and form chemical bonds. Such reactions can be expressed as a simple equation:

$$\text{Reactants} \rightarrow \text{Products}$$

All such reactions must balance, so that the total number and kind of atoms are the same on both sides. For example, oxygen and hydrogen can form water by the reaction

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

This reaction balances because each side has four hydrogen atoms and two oxygen atoms. In the process of this reaction we can observe both chemical changes (the rearrangement of atoms) and physical changes (the transformation of hydrogen and oxygen gases into liquid water with different properties).

Chemical Reactions and Energy: Rolling down the Chemical Hill

Why do chemical reactions take place at all? The fundamental reason, as so often happens with natural phenomena, has to do with energy, as described by the laws of thermodynamics (see Chapters 12 and 13).

Consider, for example, one of the electrons in the neutral sodium atom shown in Figure 23-2*a*. This electron is moving around the nucleus, so it has kinetic energy. In addition, the electron possesses potential energy because it is a certain distance from the positively charged nucleus. Finally, the electron has an additional component of potential energy because of the electrical repulsion between it and all the other electrons in the atom. This situation is analogous to the small contribution to Earth's potential energy from the gravitational attractions of the other planets. The sum of these three energies—the kinetic energy associated with the electron's motion, the potential energy associated with its attraction to the nucleus, and the potential energy associated with the other electrons—is the total energy of the single electron in its shell.

The atom's total energy is the sum of the energies of all its electrons. For the isolated sodium atom in Figure 23-2, the total energy is the sum of the energy of the 11 electrons; for the chlorine atom, it is the sum of the energies of the 17 electrons. The total energy of the sodium–chlorine system is the sum of the individual energies of the two atoms.

Now think about what happens to the energy of the sodium–chlorine system after the ionic bond has formed. The force on each electron is now different than it was before. For one thing, the number of electrons in each atom has changed; for another, the atoms are no longer isolated, so electrons and protons in the sodium atom can exert forces on electrons in the chlorine, and vice versa. Consequently, the energies of all the electrons shift a little due to the formation of the bond. This means that each electron finds itself in a slightly different position with regard to the nucleus, moving at a slightly different speed, and experiences a slightly different set of forces than it did before. The total energy of each electron is different after the bond forms, the total energy of each atom is different, and the total energy of the system is different.

Whenever two or more atoms come together to form chemical bonds, the total energy of the system is different after the bonds form than it was before. Two possibilities exist: either (1) the final energy of the two atoms is less than the initial energy or (2) the final energy is greater than the initial energy.

The reaction that produces sodium chloride from sodium and chlorine is an example of the first kind of reaction, in which the total energy of the electrons in the system is lower after the two atoms have come together. According to the first law of thermodynamics, the total energy must be conserved, and that difference in energy is given off during the reaction in the form of heat, light, and sound (there is an explosion). A chemical reaction that gives off energy in some form is said to be *exothermic*.

Many examples of exothermic reactions occur in everyday life. The energy that moves your car is given off by the explosive chemical combination of gasoline and oxygen in the car's engine when it is ignited by a spark. The chemical reactions in the battery that runs your Walkman also produce energy, although in this case some of the energy is in the form of kinetic energy of electrons in a wire. At this moment, cells in your body are breaking down molecules of the sugar glucose to supply the energy you need to live.

If the final energy of the electrons in a reaction is greater than the initial energy, then you have to supply energy to make the chemical reaction proceed. Such reactions are said to be *endothermic*. The chemical reactions that go on when you are cooking (frying an egg, for example, or baking a cake) are of this type. You can put the ingredients of a cake together and let them sit for as long as you like, but nothing happens until you turn on the oven and supply energy in the form of heat. When the energy is available, electrons can move around and rearrange their chemical bonds. The result: a cake, where before there was only a mixture of flour, sugar, and other materials.

You can think of chemical reactions as being analogous to a ball lying on the ground. If the ball happens to be at the top of a hill, it lowers its potential energy by rolling down the hill, giving up the excess energy in the form of frictional heat. If the ball is at the bottom of a hill, you have to do work on it to get it to the top. In the same way, exothermic reactions are systems that "roll down the hill," going to a state of lower energy and giving off excess energy in some form. Endothermic reactions, on the other hand, have to be "pushed up the hill," and hence absorb energy from their surroundings.

Develop Your Intuition: Hot Packs and Cold Packs

Sports trainers and coaches often carry instant cold packs or hot packs for treating minor sprains and injuries. You knead the pack with your fingers to produce the low or high temperatures and then apply the pack to the injury. How do these packs work?

Both packs contain two separate compartments before they are activated. The cold pack contains water in one compartment and crystals of the compound ammonium nitrate in the other. The hot pack has water in one compartment and crystals of magnesium sulfate or calcium chloride in the other. In both cases, kneading the pack breaks the partition between the compartments, allowing the crystals to dissolve in the water.

The interesting phenomena here are the different directions of heat transfer in these two solution processes. Ammonium nitrate absorbs heat as it dissolves; it is one of the few substances that requires energy to dissolve but does so spontaneously anyway. The result is a cold pack. On the other hand, the dissolving of magnesium sulfate or calcium chloride is an exothermic reaction and gives off heat. In both cases, once the crystals have dissolved, the packs slowly return to room temperature and do not renew their effects with further kneading. Once the crystals have dissolved, they do not return to their dry state by themselves.

A chemical cold pack absorbs heat when the inside reactants are allowed to mix together; a hot pack releases heat.

Connection
The Clotting of Blood

Whenever you get a cut that bleeds, your blood begins a remarkable and complex sequence of chemical reactions called "clotting." Normal blood is a liquid crowded with cells and chemicals that distributes nutrients and energy throughout your body. Blood flows freely through the body's circulatory system. However, when that system is breached and blood escapes, the damaged cells cause the release of a molecule called prothrombin.

Prothrombin itself is inactive, but other blood chemicals convert it into the active chemical thrombin. The thrombin reacts to break apart other normally stable chemicals that are always present in blood and thus produces small molecules that immediately begin to bind together into chains. The newly formed material, called fibrin, congeals quickly and forms a tough fiber net that traps blood cells and seals the break in minutes.

Clotting reactions differ, depending on the nature of the injury and the presence of foreign matter in the wound. Biologists have discovered more than a dozen separate chemical reactions that may occur during the process. Several diseases and afflictions may occur if some part of this complex chemical system is not functioning properly. Hemophiliacs lack one of the key clotting chemicals and so may bleed continuously from even small cuts. Some lethal snake venoms, on the other hand, work by inducing clotting in a closed circulatory system. ●

THINKING MORE ABOUT

Atoms in Combination: Life Cycle Costs

Every month, chemists around the world develop thousands of new materials and bring them to market. Some of these materials do a particular job better than those they replace, some do jobs that have never been done before, and some do jobs more cheaply. However, all of them share one property: when the useful life of the product of which they are a part is over, they will have to be disposed of in a way that is not harmful to the environment. Until very recently, engineers and planners had given little thought to this problem.

For example, think about the battery in your car. The purchase price covers the cost of mining and processing the lead in its plates, pumping and refining the oil that was made into its plastic case, assembling the final product, and so on. When that battery reaches the end of its useful life, all these materials have to be dealt with responsibly. For example, if you throw the battery into a ditch somewhere, the lead may wind up in nearby streams and wells.

One way of dealing with this sort of problem is to recycle materials—pull the lead plates out of the battery, process them, and then use them again. But even in the best system, some materials can't be recycled, either because they have become contaminated with other materials during use or because we don't have the technologies to recycle them. These materials have to be disposed of in a way that isolates them from the environment. The question becomes, "Who pays?"

Traditionally in the United States, the person who does the dumping—in effect, the last user—must see to the disposal. In some European countries, however, a new approach is being introduced. Called "life cycle costing," this approach is built around the proposition that once a manufacturer uses a material, he or she owns it forever and is responsible for its disposal. The cost of a product such as a new car, then, has to reflect the fact that someday that car may be abandoned and the manufacturer will have to pay for its disposal.

Life cycle costing increases the price of commodities, contributing to inflation in the process. What do you think the proper trade-off is in this situation? How much extra cost should be imposed up front compared to eventual costs of disposal?

Summary

Atoms link together by **chemical bonds,** which form when a rearrangement of electrons lowers the potential energy of the electron system, particularly by the filling of outer electron shells. **Ionic bonds** lower chemical potential energy by the transfer of one or more electrons to create atoms with filled shells. The positive and negative ions created in the process bond together through electrostatic forces. In metals, on the other hand, isolated electrons in the outermost shell wander freely throughout the material and create **metallic bonds. Covalent bonds** occur when adjacent atoms or molecules share bonding electrons. Hydrogen bonding and van der Waals forces are special cases involving the distortion of electron distributions to create electrical polarity—regions of slightly positive and negative charge that can bind together.

All materials, from building supplies and fabrics to electronic components and food, have properties that arise from the kinds of constituent atoms and the ways those atoms are bonded together. The high **strength** of materials such as stone and synthetic fibers relies on interconnected networks of ionic or covalent bonds, while many soft and pliable materials such as wax and graphite incorporate weak van der Waals forces. Composite materials, such as plywood, fiberglass, and reinforced concrete, merge the special strengths of two or more materials.

Chemical bonds break and form during **chemical reactions,** which may involve the synthesis or decomposition of chemical elements or compounds.

Key Terms

chemical bonds The forces that hold atoms together in stable configurations to form molecules. (p. 494)

chemical reaction The formation or breaking apart of chemical bonds. (p. 506)

covalent bond Chemical bond formed when two or more atoms in a molecule share electrons. (p. 497)

ionic bond Chemical bond formed when one atom gives up one or more electrons to another atom, creating an electrical attraction between the atoms. (p. 494)

metallic bond Chemical bond formed when many atoms share the same electrons. (p. 496)

strength The ability of a solid to resist changes in shape. (p. 502)

Review

1. The basic physical properties of all materials depend on three essential features. What are they?
2. Which is more important in explaining how and why elements combine, the nucleus or the electrons?
3. What is a valence electron? Why are valence electrons so important in understanding how elements combine?
4. What are inert gases? How many valence electrons does an inert gas have, and why does this make them so chemically unreactive?
5. Does an element with an outer shell full of valence electrons have a higher or lower energy state than an element with an unfilled outer shell? How does this affect its stability, and what is the significance of this?
6. What are three simple strategies atoms adopt to achieve a full valence shell of electrons?
7. What are three types of chemical bonds? What do they all share in common?
8. Describe the ionic bond. Give an example of a material that results from such a bond.
9. Is an ionic bond generally strong? What force holds it together?
10. Do you think the compound sodium chloride (table salt) is lower in energy as a compound than as two separate elements? Explain.
11. Can an ionic bond involve more than a single electron transfer? Explain.
12. Why are materials composed of ionic bonds often brittle?
13. Describe the metallic bond. How is it similar to an ionic bond? How does it differ?
14. One description of a metal is a collection of positive ions held together in a crystal structure by the glue of free-floating electrons. What properties of metal follow from the structure of the metallic bond? Why can you hammer or pull a metal without it shattering or breaking easily?
15. Describe the covalent bond. How does this bond differ from an ionic bond?
16. What types of bonds do carbon atoms form?
17. The study of the chemistry of carbon has its own name. What is this name, and why is so much attention given to the study of this one particular element?
18. What is a single bond? A double bond?
19. What is polarization? What is a polar molecule?

20. Describe the hydrogen bond. How does it differ from ionic and covalent bonds?
21. How does the polar nature of water make it an effective dissolving agent?
22. Are all hydrogen atoms involved in hydrogen bonds?
23. Describe the van der Waals force.
24. How are van der Waals forces like hydrogen bonds? How are they different?
25. Graphite and diamond are both made of carbon. How do they differ in structure, and what properties do they exhibit as a result of this difference?
26. What determines the strength of a material? Which types of chemical bonds are strongest?
27. What are the three types of strength of materials recognized by scientists and engineers? Are these types of strength independent of one another?
28. What is the elastic limit of a material?
29. What is a composite material? What are the benefits of such materials?
30. What is the difference between a composite material and a compound? Give an example of each.
31. Identify objects around you that use the three kinds of chemical bonding or the three kinds of attraction discussed in this chapter. Which objects incorporate two or more kinds of these bonds or attractions?
32. Chemical reactions involve the rearrangement of the atoms in compounds and elements. What does it mean to say that a reaction must be balanced?
33. Why do chemical reactions occur? What is the role of energy in making them happen?
34. What is an exothermic reaction? Is heat given off or absorbed in this type of reaction?
35. What is an endothermic reaction? Is heat given off or absorbed in this type of reaction?
36. How is the energy change in a chemical reaction analogous to a ball on the ground in a hilly area?

Questions

1. Is hydrogen or helium a better choice to fill a balloonlike airship, or blimp? Compare the advantages and disadvantages of each gas, keeping in mind the properties that flow from their respective atomic structures.
2. What types of chemical bonds are the strongest and why?
3. Why are covalent bonds so prevalent in biological molecules?
4. If you had your choice, would you build a house with ionic-bonded materials or with materials held together by van der Waals forces?
5. Do the properties of a newly formed chemical compound tend to differ from or be the same as the properties of the individual elements that compose it? Give several examples to support your answer. Is your answer different in the case of an alloy?
6. Diamonds and graphite are both made from carbon atoms. Why is graphite so much weaker?
7. An ionic bond is formed when an electron from one atom is transferred to another atom. Explain why NaF (sodium fluoride) forms an ionic bond.
8. Explain why NaCl (sodium chloride) forms an ionic bond. Why do magnesium and chlorine form $MgCl_2$ more readily than MgCl?
9. A molecule of methane (CH_4) has a carbon atom that is bonded to four hydrogen atoms. The C—H bonds are covalent, meaning that the carbon atom shares an electron with each hydrogen atom. Explain why methane is a good candidate for a covalently bonded molecule.
10. Write the chemical formula for the following covalent compounds: carbon chloride (carbon and chlorine) and hydrogen chloride (hydrogen and chlorine).
11. Write the chemical formula for the ionic compound magnesium oxide (magnesium and oxygen).
12. Potassium iodide can be used as a thyroid-blocking agent in the event of a radiation emergency. Write the chemical formula for the ionic compound potassium iodide (potassium and iodine).
13. Magnesium (Mg) and bromine (Br) form an ionic compound. What is its chemical formula? Which element becomes the positive ion in this compound?
14. Lead (Pb) and sulfur (S) combine to form an ionic compound. What is its chemical formula? Which element is the positive ion in this compound?
15. Aluminum (Al) and chlorine (Cl) combine to form an ionic compound. What is its chemical formula? Which element is the positive ion in this compound?
16. Explain why chlorine is much more reactive than argon, even though they appear in adjacent spots on the periodic table.

Problems

1. Some elements readily form ionic or covalent bonds and some elements do not participate in chemical bonding at all. In the table below, identify each element by its atomic number (see the periodic table in Appendix C), indicate (with a Yes or No) whether or not the element is likely to participate in chemical bonding, and give your reasoning.

Atomic Number	Element Name	Chemical Bonding	Reason
1			
2			
6			
7			
8			
10			
11			
13			
14			
16			
17			
18			
19			
20			
26			
29			
47			
79			
82			
86			

Investigations

1. What materials were used for the construction of buildings, furniture, and transportation devices in the United States 200 years ago? What modern technologies would be difficult or impossible if we could only use those materials?
2. Gold is often sold in a less than pure form as an alloy. What does it mean to say that you are buying, say, a necklace that is 14-carat gold? What is a carat and what other material(s) is gold alloyed with and why? Similarly, what is sterling silver?
3. Why is fluorine so effective in helping to prevent tooth decay? Investigate how fluorine acts to do this and examine the history of its use, including its controversial introduction into municipal water systems that started in the late 1940s. Over the years, what concerned opponents? Did the concerns differ according to whether the critic was conservative or liberal?
4. Dissect a disposable diaper. How many kinds of materials can you identify? What are the key properties of each? What kind of chemical bonding might contribute to the distinctive properties of these materials? Investigate the arguments for and against using disposable diapers.
5. Investigate how steel is made. Follow the process of production from extraction of the ore from the ground, to the separation of the iron from the ore, to the steel mill, and through the finishing processes. What are some of the different types of steel produced? What are some of the additives used in various alloys? Pay particular attention to the properties of the materials involved at each stage of the process as you explore this broad topic.
6. Visit a sports equipment store. Learn about the new materials that are used in tennis rackets, football helmets, and sports clothing.
7. Write a short story in which a new material with unique properties plays a central role.
8. What kinds of materials do surgeons use to replace broken hip bones? What are the advantages of this material?

WWW Resources

See the *Physics Matters* home page at **www.wiley.com/college/trefil** for valuable web links.

1. **http://www.sfu.ca/person/lower/TUTORIALS/chembond/index.html** The "All about chemical bonding" tutorial by Stephen Lower of Simon Fraser University.

2. **http://www.chemistry.org/portal/a/c/s/1/acsdisplay.html?DOC=sitetools\periodic_table.html** A sophisticated periodic table of the elements site by the American Chemical Society. See especially the filling of electron shells.
3. **http://educ.queensu.ca/~science/main/concept/chem/c07/C07TPSU3.html** A discussion conceptualizing kinds of bonds as everyday social situations.
4. **http://www.kathysnostalgiabilia.com/lemelson.htm** The history of nylon and its impact on the U.S. chemical industry and culture.
5. **http://physics.uwstout.edu/strength/indexfbt.htm** A complete online course devoted to the strength of materials.
6. **http://micro.magnet.fsu.edu/electromag/java/atomicorbitals/index.html** *Atomic Orbitals.* Short and sweet visualization of electronic orbital shapes from Florida State University.