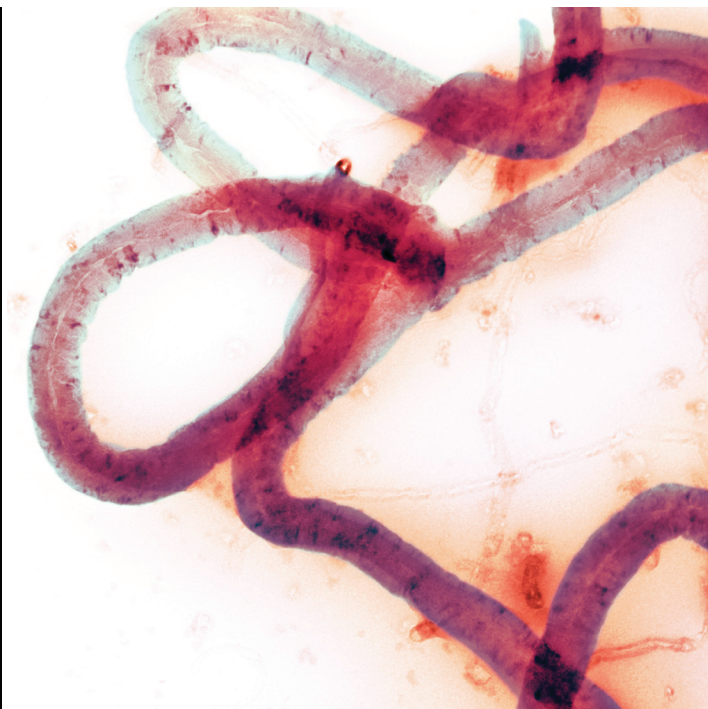
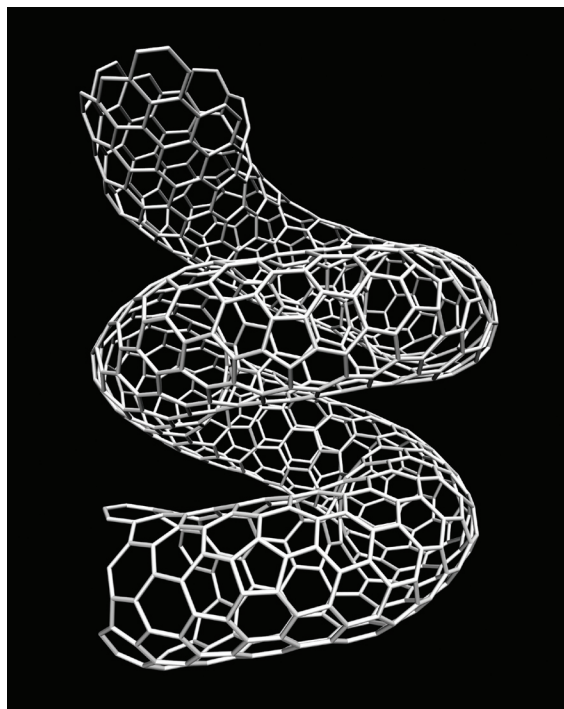


## INTERCHAPTER M

# Carbon and Silicon



Carbon nanotubes have potential applications ranging from new composites for use in ultralight automobiles and bullet-proof clothing to ultrathin wires. The discovery of carbon nanotubes helped launch the newly emerging field of nanotechnology.

Carbon is widely distributed in nature both as the free element and in compounds. The great majority of naturally occurring carbon is found in coal, petroleum, **limestone**,  $\text{CaCO}_3(s)$ , **dolomite**,  $\text{MgCa}(\text{CO}_3)_2(s)$ , and a few other deposits. Carbon is also a principal element in all living matter, and the study of its compounds forms the vast fields of organic chemistry and biochemistry. Carbon is decidedly nonmetallic, whereas silicon is a semimetal.

Silicon constitutes 28% of the mass of the earth's mantle and is the second most abundant element in the mantle, exceeded only by oxygen. Silicon does not occur as the free element in nature; it occurs primarily as silicon dioxide,  $\text{SiO}_2(s)$  and in numerous silicates. Most sands are primarily silicon dioxide, and many rocks and minerals are silicate materials. The major use of silicon is in the manufacture of glass and transistors. Pure silicon is an insulator, but it can be converted to a semiconductor by addition of trace amounts of certain elements.

The physical properties of carbon and silicon are given in Table M.1, and major sources and uses of these elements are given in Table M.2.

### M-1. Diamond and Graphite Are Allotropic Forms of Carbon

As discussed in Section 15-10, the two most important allotropic forms of solid carbon are diamond and graphite. Recall that diamond has an extended, covalently bonded tetrahedral structure and that graphite has a layered structure. Because diamond is the hardest naturally occurring mineral known, it is extensively used as an abrasive and cutting material where a very high resistance to wear is required. Graphite, however, is the stable form of carbon at ordinary temperatures and pressures. It exists as a solid at a higher temperature than any other material known ( $3825^\circ\text{C}$ ).

To produce diamonds (Figure M.1), it is necessary to subject graphite to a pressure and temperature that lie above the graphite-diamond equilibrium line in Figure M.2. For example, at  $25^\circ\text{C}$ , a pressure greater than 15000 bar is required. High pressure favors the solid form with the greater density (smaller molar volume). However, even at  $25^\circ\text{C}$  and 15000 bar, the rate of conversion of graphite to diamond is extremely slow.

TABLE M.1 Physical properties of carbon and silicon

Element	Symbol	Atomic mass	Melting point/ $^\circ\text{C}$	Boiling point/ $^\circ\text{C}$	Density at $25^\circ\text{C}/\text{g}\cdot\text{cm}^{-3}$
carbon	C	12.0107	sublimes at 3825		2.2 (graphite), 3.513 (diamond)
silicon	Si	28.0855	1414	3265	2.3296

TABLE M.2 Major sources and uses of carbon and silicon

Element	Sources	Form	Uses
carbon	coal and petroleum	coke	fuels; production of iron and steel, furnace linings, electrodes
		graphite	lubricant, fibers, pencils, airframe structures
		activated charcoal	decolorizer, air purification, catalyst
		carbon black	rubber and printing inks
		diamond	drill bits, abrasives, jewelry
silicon	quartzite or sand, $\text{SiO}_2(s)$		steel alloys; silicones; semiconductor in integrated circuits, rectifiers, transistors, solar cells

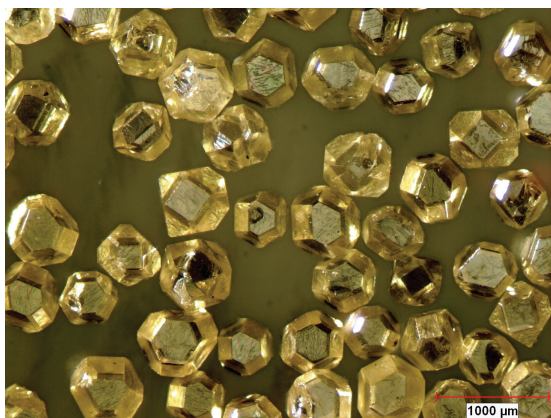


Figure M.1 Synthetic diamonds.

The first successful synthesis of diamonds from graphite was carried out at General Electric Laboratories in 1955. At 2200°C and 150 000 bar, essentially complete conversion of graphite to diamond occurs in a few minutes. A rapid decrease in the pressure and temperature traps the carbon in the diamond form.

Another important form of carbon is **activated carbon**, an amorphous form characterized by its ability to absorb large quantities of gases (Figure M.3). Activated carbon can be produced by heating wood or other carbonaceous material to 800–900°C with

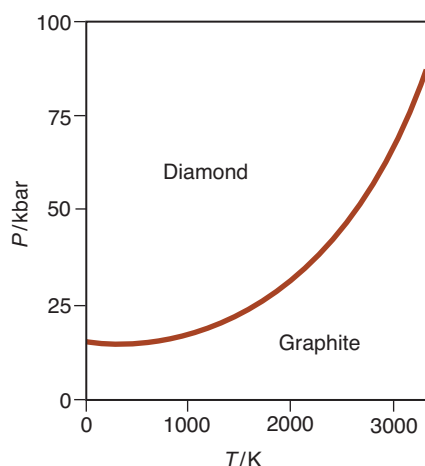


Figure M.2 Graphite-diamond phase diagram. Above the curve, diamond is the stable form of solid carbon; below the curve, graphite is the stable form. At 25°C the equilibrium transition pressure for the conversion of graphite to diamond is 15 000 bar. Diamond has a higher density than graphite, so diamond is the stable form of C(s) at high pressures. Diamonds at one bar are unstable with respect to graphite, but the rate of conversion of diamonds to graphite at 25°C is negligible.



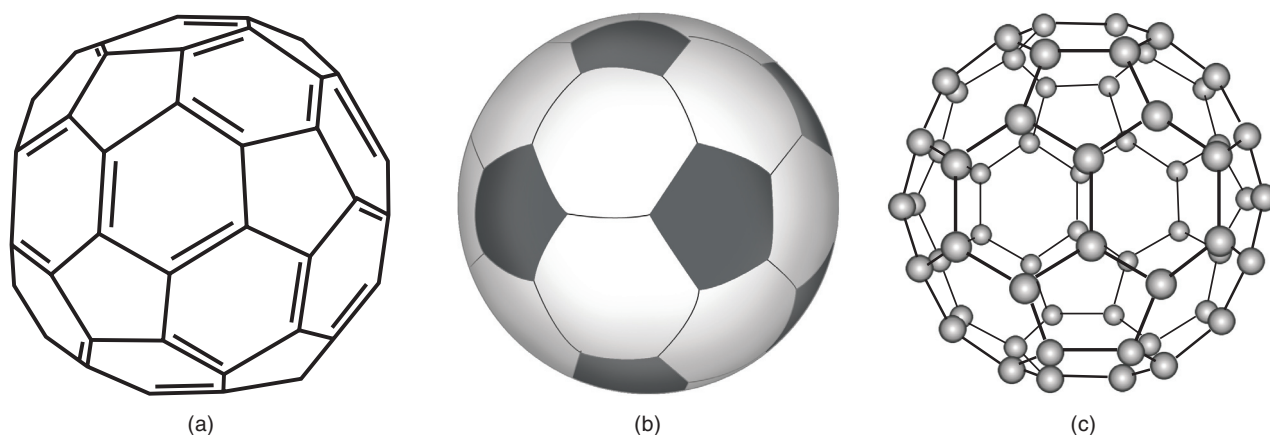
Figure M.3 The flask on the left contains nitrogen dioxide,  $\text{NO}_2(g)$ , and the dish in front contains activated carbon, a black powder. The flask on the right shows the result of adding the activated carbon to the flask containing  $\text{NO}_2(g)$ . Observe that the  $\text{NO}_2(g)$  has been completely absorbed.

steam or carbon dioxide. This process results in a porous product with a honeycomb structure. The ability of activated carbon to absorb large quantities of gases is due to its very large surface area, which averages about 1000 square meters per gram.

## M-2. Fullerenes Are Another Allotrope of Carbon

In 1985 Harold Kroto, Robert Curl, and Richard Smalley, together with graduate students Jim Heath and Sean O'Brien, were attempting to create long carbon chains similar to those observed by radioastronomers near red dwarf stars. In their experiment they used a laser to vaporize a graphite disk, analyzing the products formed using a mass spectrometer. To their surprise, in addition to the sought-after compounds, they observed a particularly stable peak at a mass of 720 u, corresponding to a formula of  $\text{C}_{60}$ . Neither the bonding in graphite nor that in diamond could account for a stable structure containing 60 carbon atoms. Although there wasn't enough product available to perform a chemical analysis, the team suggested that the molecule was a **truncated icosahedron**, a three-dimensional structure composed of 12 pentagons and 20 hexagons shaped like a soccer ball (Figure M.4)—a never-before-observed allotrope of carbon. The team named the new molecules **buckminsterfullerenes**, after architect Richard Buckminster Fuller, who designed the geodesic dome. They are also often called by their more colloquial





**Figure M.4** The truncated icosahedron structure of  $C_{60}(s)$ , buckminsterfullerene (a), is similar to that of a soccer ball (b). Each carbon atom in buckminsterfullerene is covalently bonded to three other atoms through a network of  $sp^2$ -hybridized orbitals (c). To form the  $C_{60}(s)$  cage, each of these orbitals is distorted from its ideal 120-degree bond angle down to 108 degrees. The  $\pi$ -electrons in the remaining  $p$ -orbitals are delocalized over the entire molecule, giving the  $C_{60}(s)$  molecule its enormous resonance stability.

name, **buckyballs**. Since their initial discovery, scientists have learned how to produce buckminsterfullerene in macroscopic quantities. Kroto, Curl, and Smalley shared the 1996 Nobel Prize in Chemistry for their discovery of fullerenes, a new allotrope of carbon (sidebar).

Since this discovery, a number of other interesting stable compounds of similar structure, collectively called **fullerenes**, have been discovered, including  $C_{70}$ ,  $C_{76}$ ,  $C_{84}$ , and so forth up to giant fullerene molecules containing more than half a million atoms. Fullerenes have been found to be a component of

**BUCKMINSTERFULLERENE DISCOVERERS** This photograph shows the team of researchers that discovered buckminsterfullerene,  $C_{60}(s)$ , a new form of carbon. Kneeling in the photo from left to right are Sean O'Brien, Rick Smalley, Harry Kroto, and Jim Heath; Bob Curl is standing in the back. Smalley, Kroto, and Curl were professors of chemistry and O'Brien and Heath were graduate students at the time (1985). Smalley, Kroto, and Curl shared the Nobel Prize in Chemistry in 1996 "for their discovery of fullerenes." Richard E. Smalley was born in Akron, Ohio, in 1943, and received his Ph.D. in chemistry from Princeton University in 1973. In 1976, he moved to Rice University. There he helped create the Rice Center for Nanoscience and Technology. He died of leukemia in 2005. Robert F. Curl, Jr. was born in Alice, Texas, in 1933, and received his Ph.D. in chemistry from the University of California at Berkeley in 1957. He joined the faculty at Rice University in 1958, where he remains today. Curl's research interests are in laser spectroscopy and chemical kinetics. Sir Harold Kroto was born in Wisbech, England, in 1939, and received his Ph.D. in chemistry from the University of Sheffield in 1964. After postdoctoral research in Canada and the United States, he returned to England to the University of Sussex in Brighton. He was knighted in 1996, and in 2005 the Kroto Research Center for Nanoscience and Technology, a multidisciplinary research center that he helped support, opened at the University of Sussex.



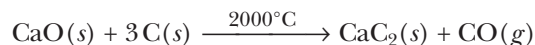
soot, observed spectroscopically in interstellar space, discovered in rocks hit by lightning, and found in meteors. One of the more intriguing findings was that of helium atoms with isotopic ratios not found on earth trapped inside fullerene cages isolated from meteorites, suggesting a possible origin outside our solar system.

Fullerenes are not the only new allotrope of carbon to form upon the vaporization of graphite; another cylindrical structure called **carbon nanotubes** has also been discovered and is the focus of much attention. The structure of carbon nanotubes is a chicken-wire-like mesh of carbon atoms composed of hexagons wrapped into a long cylinder 1–2 nm in diameter and up to a millimeter or longer in length (Frontispiece). Nanotubes may eventually find uses as new building materials for cars and aircraft, ultra-strong fibers for clothing, ultrathin flexible wires, and a variety of other innovative applications. The discovery of nanotubes has fueled the growth of a new field of research called **nanotechnology**, the goal of which is the production of molecular-sized machines and electronic devices that are several to a hundred nanometers in size.

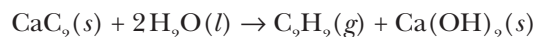
### M-3. Carbon Ranks Second among the Elements in the Number of Compounds Formed

At one time, chemists divided all compounds into two classes: **inorganic** and **organic**. Inorganic compounds were classified as those that could be obtained from minerals and other inanimate sources, whereas organic compounds were those obtained from living or once-living sources. Today chemists still classify compounds as either inorganic or organic, but now organic compounds are almost always considered to be simply compounds consisting primarily of carbon atoms. Carbon atoms are unusual in that they tend to form long covalently bonded chains and rings. The vast majority of all known compounds, and particularly almost all biologically important compounds, are organic compounds. In spite of the definition of organic compounds given above, there are a number of simple carbon compounds that chemists consider to be inorganic compounds. We shall discuss some of these inorganic compounds in this section (various organic compounds are discussed in Interchapters F, G, H, P, and R).

Binary compounds in which carbon is combined with less electronegative elements are called **carbides**. One of the most important carbides is calcium carbide. It is produced industrially by the reaction of lime,  $\text{CaO}(s)$ , and coke (a form of solid carbon that can be derived from coal) according to



Calcium carbide is a gray-black, hard solid with a melting point over  $2000^\circ\text{C}$ . Calcium carbide reacts with water to produce acetylene. The equation for the reaction is



At one time, this reaction represented one of the major sources of acetylene for the chemical industry and for oxyacetylene welding, but its use has declined because of the high cost of the energy used in the production of the calcium carbide. Acetylene is now produced from petroleum and natural gas.

One other industrially important carbide is silicon carbide,  $\text{SiC}(s)$ , also known as **carborundum**. Carborundum is one of the hardest known materials and is used as an abrasive for cutting metals and polishing glass. Its structure is similar to the cubic crystal structure of diamond.

Carbon has a number of oxides, but only two of them are particularly stable. When carbon is burned in a limited amount of oxygen, carbon monoxide predominates. When an excess of oxygen is used, carbon dioxide results. Carbon monoxide,  $\text{CO}(g)$ , is an odorless, colorless, tasteless gas that burns in oxygen to produce carbon dioxide. Carbon monoxide is highly poisonous because it binds to hemoglobin much more strongly than oxygen does and thereby blocks the oxygen-carrying ability of hemoglobin. Carbon monoxide is used as a fuel and as a reducing agent in metallurgy.

Carbon dioxide,  $\text{CO}_2(g)$ , is an odorless, colorless gas. Over 50% of the carbon dioxide produced industrially is used as a refrigerant, either as a liquid or as a solid (dry ice), and about 25% is used to carbonate soft drinks. Carbon dioxide is also used in household fire extinguishers. The phase diagram of carbon dioxide is discussed in Section 15-8.

Carbon forms several sulfides, but only one of them, carbon disulfide,  $\text{CS}_2(l)$ , is stable at room tem-

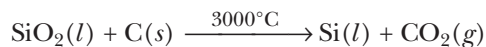
perature. Carbon disulfide is a colorless, poisonous, flammable liquid. The purified liquid has a sweet, pleasing odor, but the commonly occurring commercial and reagent grades have an extremely disagreeable odor that is due to organic impurities. Large quantities of  $\text{CS}_2(l)$  are used in the manufacture of rayon, carbon tetrachloride, and cellophane. It is also used as a solvent for a number of substances.

Carbon also forms several important nitrogen-containing compounds. Hydrogen cyanide,  $\text{HCN}(g)$ , is a colorless, extremely poisonous gas that dissolves in water to form the very weak acid, hydrocyanic acid. Salts of hydrocyanic acid, called cyanides, are prepared by direct neutralization. Sodium cyanide,  $\text{NaCN}(s)$ , is used in the extraction of gold and silver from their ores and in the electroplating industry.

#### M-4. Silicon Is a Semimetal

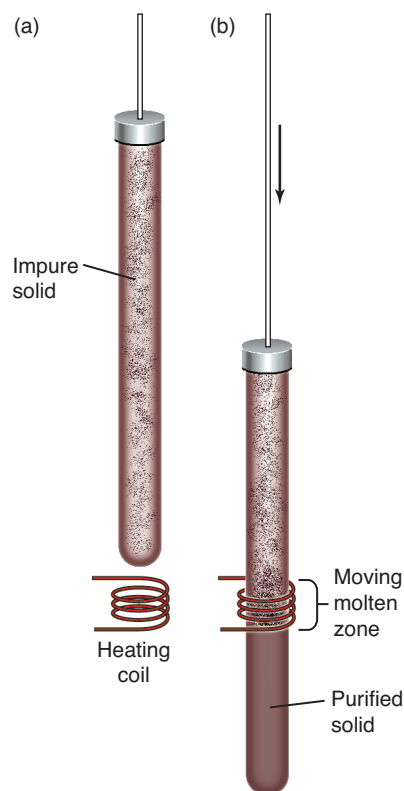
Silicon, the most important industrial semimetal, has a gray, metallic luster. Its major use is in the manufacture of transistors and other electronic devices. **Amorphous silicon** is a solid form of silicon that does not have a crystalline structure. It is used in the fabrication of some solar energy devices because of its relatively low cost.

Elemental silicon is made by the high-temperature reduction of silicon dioxide (the major constituent of numerous sands) with carbon according to



The 98% pure silicon prepared by this reaction can be further purified by a special method of recrystallization called **zone refining**. In this process, solid silicon is packed in a tube that is mounted in a vertical position with an electric heating loop around the base (Figure M.5). The solid near the heating loop is melted by passing a current through the loop, and the tube is then lowered very slowly through the loop. As the melted solid cools slowly in the region of the tube below the heating loop, pure crystals separate out, leaving most of the impurities in the moving molten zone. The process can be repeated as often as necessary to achieve purities up to 99.9999%.

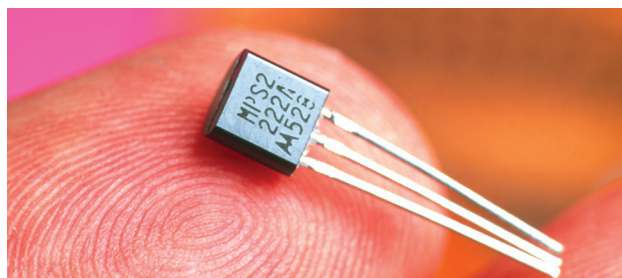
Recall from Section 15-11 that silicon is a semiconductor. We learned there that the energy gap between the conduction band and the valence band is comparable to thermal energies at room temper-



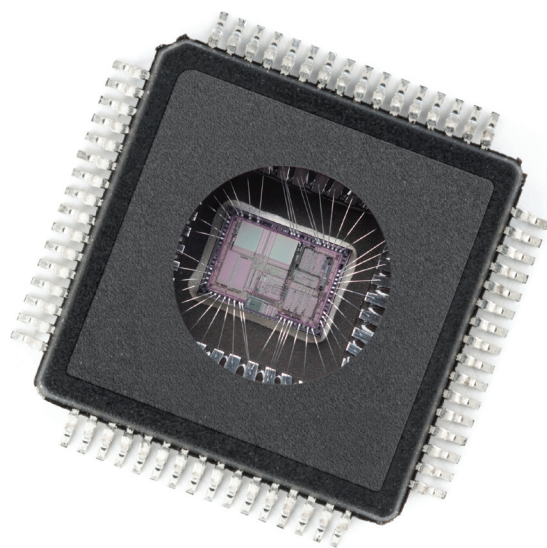
**Figure M.5** Zone refining. An impure solid is packed tightly in a glass tube, and the tube is lowered slowly through a heating coil that melts the solid. Pure solid crystallizes out from the bottom of the melted zone, and the impurities concentrate in the moving molten zone.

ature in a semiconductor. Therefore, some of the valence electrons will be thermally excited into the conduction band. Figure 15.37 illustrates the difference between a metal, an insulator, and a semiconductor. Transistors are possible because the electrical conductivity of a semiconductor like silicon can be controlled (Figure M.6).

Extremely pure silicon is required for manufacturing semiconductors because the electrical conductivity of silicon can be increased by the addition of selected impurity atoms to ultrapure silicon (Figure M.7a). For example, an ***n*-type** (*n* for negative) silicon semiconductor is produced when trace amounts of atoms with five valence electrons, such as phosphorus or antimony, are added to silicon, which has four valence electrons. The excess valence electrons on the impurity atoms, which substitute for some of the silicon atoms in the crystal, become the current carriers in the crystal (Figure M.7b). A ***p*-type** (*p* for positive) semiconductor is produced when trace



(a)

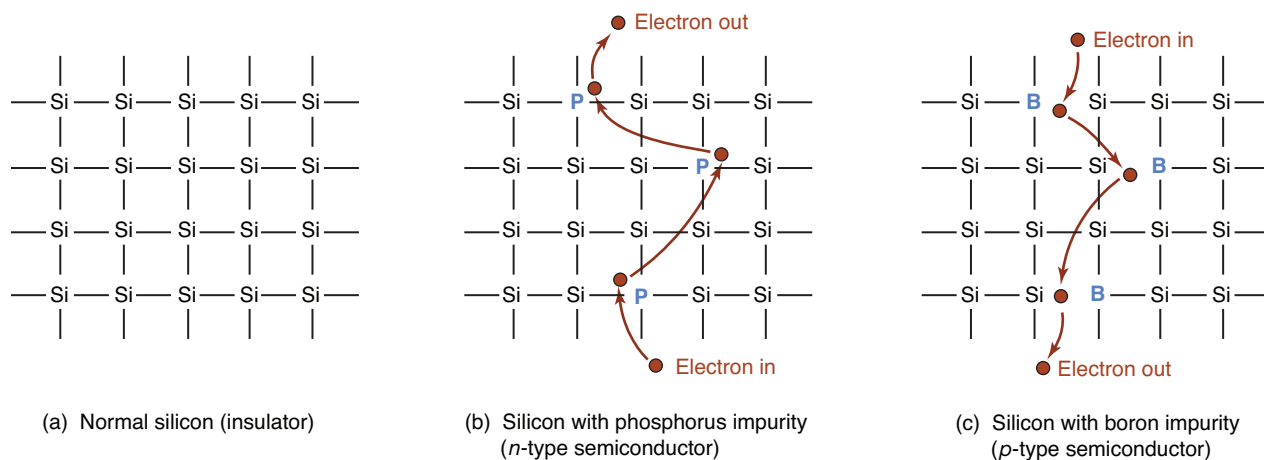


(b)

**Figure M.6** (a) A single transistor. Transistors are electronic devices used to switch or amplify signals. (b) An integrated circuit such as this computer processing chip can contain over a million transistors per square millimeter. Computing power has increased dramatically over the past few decades because of the tremendous increase in the number of transistors that can be made to fit inside a single chip.

amounts of atoms with three valence electrons, such as boron or indium, are added to silicon. The deficiency of valence electrons on the impurity atoms functions as “holes” by means of which electrons can “hop” through the silicon crystal (Figure M.7c). By adding precise amounts of impurities of carefully controlled composition to crystals of extremely pure silicon, you can obtain the desired electrical properties.

It would be difficult to exaggerate the impact of semiconductor devices on modern technology. With their minute size and very low power requirements, these devices have made possible computers with incredible computing, storage, and retrieval capabilities. It is possible to make a computer memory chip with over a million transistors in a space of only a square millimeter (Figure M.6b).



**Figure M.7** Comparison of normal, *n*-type, and *p*-type silicon. (a) Silicon has four valence electrons, so each silicon atom forms four 2-electron bonds to other silicon atoms. (b) Phosphorus has five valence electrons; thus, when a phosphorus atom substitutes for a silicon atom in a silicon crystal, there is an unused valence electron on each phosphorus atom that can become a conduction electron. (c) Boron has only three valence electrons; thus, when a boron atom substitutes for a silicon atom in a silicon crystal, an electron vacancy (a “hole”) results. Electrons from the silicon valence bond can move through the crystal by hopping from one vacancy site to another.



### M-5. The Most Common and Important Compounds of Silicon Involve Oxygen

Elemental silicon is fairly unreactive; it does not react with most acids. Silanes, the silicon hydrides analogous to the hydrocarbons, are much less stable than the corresponding hydrocarbons. Except for monosilane,  $\text{SiH}_4(g)$ , and disilane,  $\text{Si}_2\text{H}_6(g)$ , they decompose spontaneously at room temperature. There are no silicon analogs of ethylene,  $\text{C}_2\text{H}_4(g)$ , or acetylene,  $\text{C}_2\text{H}_2(g)$ , but a few compounds with silicon-silicon double bonds are known.

Silicon does react at ordinary temperatures with the halogens to give tetrahalides. It also reacts with dilute alkaline solutions to give silicates, which are found in numerous minerals and in asbestos, mica, and clays. Cement, bricks, tiles, porcelains, glass, and pottery are all made from silicates. All silicates involve silicon-oxygen single bonds.

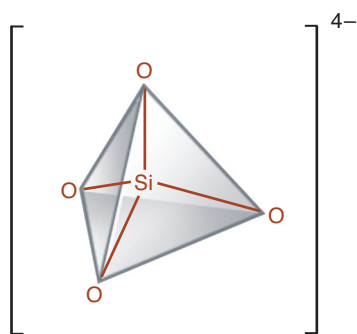


Figure M.8 The orthosilicate ion,  $\text{SiO}_4^{4-}$ .

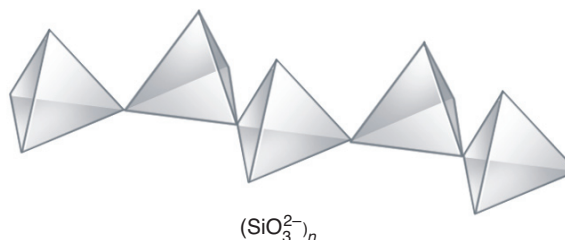
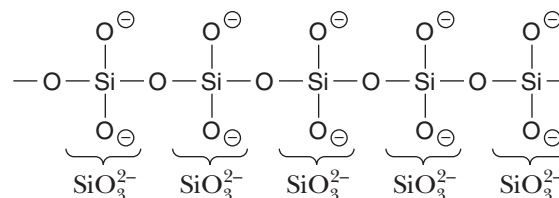


Figure M.10 Tetrahedral  $\text{SiO}_4^{4-}$  units are linked together through oxygen atoms that are shared by tetrahedra to form straight-chain silicate polyanions.

The simplest silicate anion is the tetrahedral **orthosilicate ion**,  $\text{SiO}_4^{4-}$  (Figure M.8). The  $\text{SiO}_4^{4-}$  ion is found in the minerals zircon,  $\text{ZrSiO}_4(s)$ , willemite,  $\text{Zn}_2\text{SiO}_4(s)$ , and sodium silicate,  $\text{Na}_2\text{SiO}_4(s)$  (Figure M.9). The minerals enstatite,  $\text{MgSiO}_3(s)$ , and spodumene,  $\text{LiAl}(\text{SiO}_3)_2(s)$ , are silicates that contain long, straight-chain silicate polyanions involving the  $\text{SiO}_3^{2-}$  chain unit (Figure M.9 and Figure M.10):

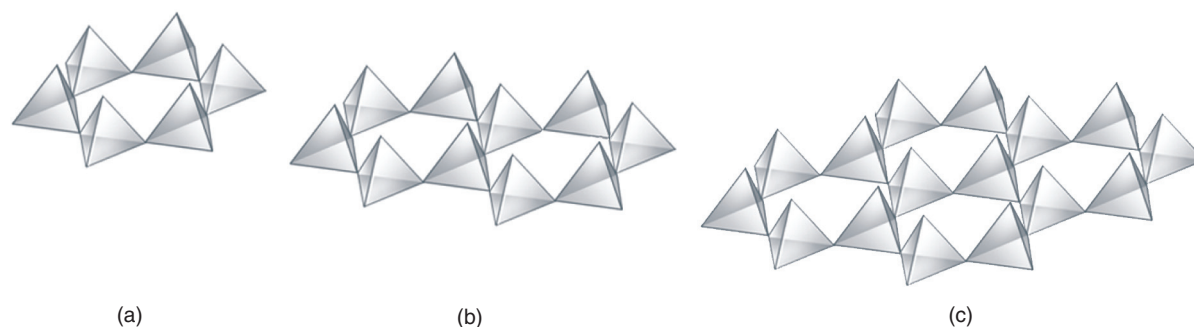


Structures that result from joining many smaller units together are called **polymers** (see Interchapters S and T). The straight-chain silicate anions shown in the preceding structure are called silicate polyanions



Figure M.9 The minerals zircon (*left*), willemite (*rear*), enstatite (*right*), and spodumene (*front*).





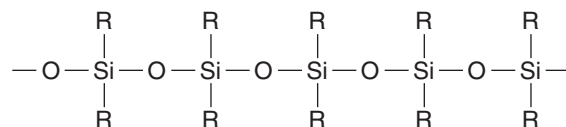
**Figure M.11** (a) The cyclic polysilicate ion  $\text{Si}_6\text{O}_{18}^{12-}$ , which occurs in the mineral beryl. Six  $\text{SiO}_4^{4-}$  tetrahedral units are joined in a ring with the tetrahedra linked by shared oxygen atoms. (b) The cyclic polysilicate ion  $\text{Si}_6\text{O}_{18}^{12-}$  can form a polymeric cyclic network like that shown here. The composition of the cyclic network is  $(\text{Si}_4\text{O}_{11}^{6-})_n$ . Asbestos has this structure. (c) Structure of polysilicate sheets composed of  $(\text{Si}_2\text{O}_5^{2-})_n$  subunits. Mica has this structure.

because they consist of many linked silicate anions. The mineral beryl,  $\text{Be}_3\text{Al}_2\text{Si}_6\text{O}_{18}(s)$ , contains the cyclic polysilicate anion  $\text{Si}_6\text{O}_{18}^{12-}$  (Figure M.11a). These cyclic polysilicate anions can themselves be joined together to form polymeric, cyclic polysilicate anions with the composition  $(\text{Si}_4\text{O}_{11}^{6-})_n$  and the structure shown in Figure M.11b. The best example of a mineral containing polymeric, cyclic polysilicate chains is **asbestos**. The fibrous character of asbestos is a direct consequence of the molecular structure of the  $(\text{Si}_4\text{O}_{11}^{6-})_n$  polymeric chains (Figure M.12a).

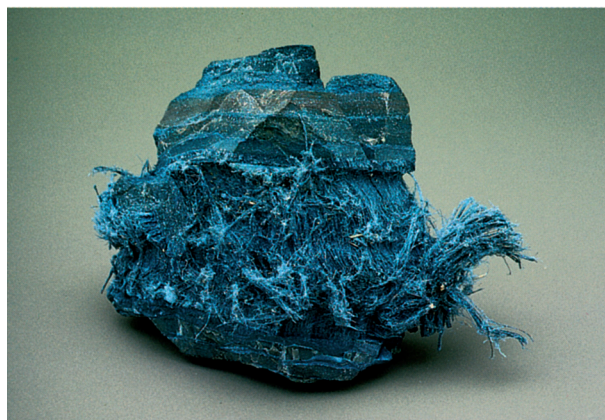
The silicate minerals **mica** and **talc** contain two-dimensional, polymeric silicate sheets with the overall silicate composition  $(\text{Si}_2\text{O}_5^{2-})_n$ . The structure of these sheets is illustrated in Figure M.11c. The ease with which mica can be separated into thin sheets and the slippery feel of talcum powder arise from the

layered structure of the silicates in these minerals (Figure M.12b).

Polysiloxanes are mixed inorganic-organic polymers involving an inorganic silicon-oxygen backbone with various organic molecules (symbolized R) covalently bound to the silicon atoms:



These polymerized siloxanes are collectively called **silicones**. By attaching different organic molecules to the backbone, various silicone compounds with a wide variety of different physical properties can be synthesized. These include liquids, gels, oils, plastics, and rubbers. Silicone compounds are used



(a)



(b)

**Figure M.12** (a) The fibrous character of asbestos is a direct consequence of its  $(\text{Si}_4\text{O}_{11}^{6-})_n$  polymeric chains. (b) The ease with which the mineral mica can be separated into thin sheets is a direct consequence of the existence of polymeric silicate sheets with the composition  $(\text{Si}_2\text{O}_5^{2-})_n$ .

TABLE M.3 Important compounds of silicon

Compound	Uses
sodium silicate, $\text{Na}_4\text{SiO}_4(s)$	soaps and detergents; silica gels; adhesives; water treatment; sizing of textiles and paper; waterproofing cement; flame retardant
silicon carbide (carborundum), $\text{SiC}(s)$	abrasive for cutting and grinding metals
silicon dioxide (silica), $\text{SiO}_2(s)$	glass manufacture; abrasives; refractory material; cement
silicones, $\left( \begin{array}{c} \text{R} \\   \\ \text{---Si---O---} \\   \\ \text{R}' \end{array} \right)_n$	lubricants; adhesives; protective coatings; coolant; fire-proofing agents; caulking; cosmetics; breast implants; gaskets; and many more applications

as caulking and sealants; as greases and synthetic lubricants; as gaskets and seals; as fire-resistant and water-proofing materials; in breast implants and cosmetic creams; and as a replacement for carcinogenic chemicals in dry cleaning—to name just a few of their applications. Some other commercially important compounds of silicon are given in Table M.3.

### M-6. Most Glasses Are Silicates

**Quartz** (Figure M.13) is a crystalline material with the composition  $\text{SiO}_2(s)$  and the crystalline structure shown in Figure M.14. When crystalline quartz

is melted and then cooled quickly to prevent the formation of crystals, a disordered three-dimensional array of polymeric chains, sheets, and other three-dimensional clusters forms. The resulting material is called quartz glass. All glass consists of such a random array.

Glass manufacturing is a 25-billion-dollar-per-year industry in the United States. The major component in glass is almost pure quartz sand. Among the other components of glass, soda,  $\text{Na}_2\text{O}(s)$ , comes from soda ash,  $\text{Na}_2\text{CO}_3(s)$ ; lime,  $\text{CaO}(s)$ , comes from limestone,  $\text{CaCO}_3(s)$ ; and aluminum oxide,  $\text{Al}_2\text{O}_3(s)$ , comes from feldspars, which have the general formula



Figure M.13 Quartz often forms large, beautiful crystals.

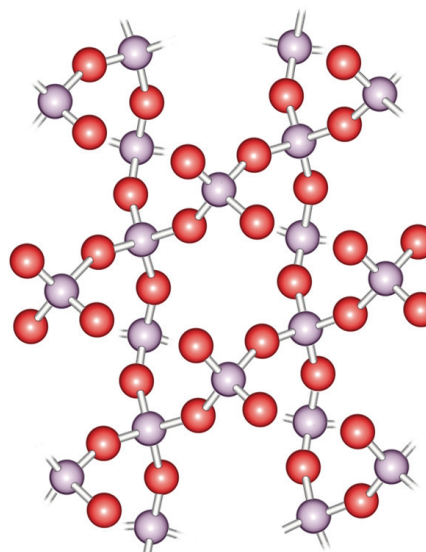
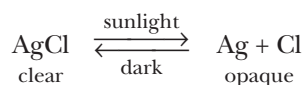


Figure M.14 The crystalline structure of quartz. Note that each silicon atom is surrounded by four oxygen atoms. The silicon atoms are linked by the oxygen atoms.

$M_2O \cdot Al_2O_3 \cdot 6SiO_2(s)$ , where M is K or Na. All the components of glass are fairly inexpensive chemicals.

A wide variety of glass properties can be produced by varying the glass composition. For example, partial replacement of  $CaO(s)$  and  $Na_2O(s)$  by  $B_2O_3(s)$  gives a glass called borosilicate glass that expands less on heating (or contracts less on cooling) and is thus used in making glass utensils meant to be heated. Colored glass is made by adding a few percent of a colored transition metal oxide, such as  $CoO(s)$  to make blue “cobalt” glass,  $Cr_2O_3(s)$  to make orange glass, and  $Au_2O_3(s)$  to make ruby glass. Because it is made with gold, ruby glass is more expensive than other colored glass (Figure M.15). Lead glass, which contains  $PbO(s)$ , has attractive optical properties and is used to make decorative, cut-glass articles.

Photochromic eyeglasses have a small amount of added silver chloride dispersed throughout and trapped in the glass. When sunlight strikes this type of glass, the tiny  $AgCl(s)$  grains decompose into opaque clusters of silver atoms and chlorine atoms according to



The chlorine atoms are trapped in the crystal lattice, so the silver and chlorine atoms can recombine in the



**Figure M.15** A decorative container made of ruby glass. The rich red color of ruby glass is due to a dispersion of very finely divided gold throughout the glass.

dark to form silver chloride, a reaction that causes the glass to become clear.

Porcelain has a much higher percentage of  $Al_2O_3(s)$  than glass and as a result is a heterogeneous substance. Porcelain is stronger than glass because of this heterogeneity and is also more chemically resistant than glass. Earthenware is similar in composition to porcelain but is more porous because it is fired at a lower temperature.

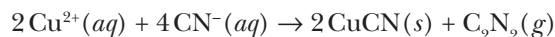
## TERMS YOU SHOULD KNOW

limestone *M1*  
 dolomite *M1*  
 activated carbon *M2*  
 truncated icosahedron *M2*  
 buckminsterfullerene,  $C_{60}(s)$  *M2*  
 buckyballs *M3*  
 fullerenes *M3*  
 carbon nanotubes *M4*  
 nanotechnology *M4*  
 inorganic *M4*  
 organic *M4*  
 carbides *M4*  
 carborundum *M4*  
 amorphous silicon *M5*  
 zone refining *M5*  
*n*-type semiconductor *M5*  
*p*-type semiconductor *M5*  
 orthosilicate ion,  $SiO_4^{4-}$  *M7*  
 polymer *M7*  
 asbestos *M8*  
 mica *M8*  
 talc *M8*  
 silicones *M8*  
 quartz *M9*

## QUESTIONS

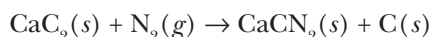
**M-1.** List the key distinctions between metals, semi-metals, and nonmetals.

**M-2.** Cyanogen,  $C_2N_2(g)$ , is a very poisonous gas that can be made by mixing warm concentrated solutions of potassium cyanide and copper(II) sulfate. The net ionic equation is:



Write the Lewis formula of cyanogen.

**M-3.** Finely powdered calcium carbide will react with nitrogen at high temperature and about two bars pressure to form calcium cyanamide,  $CaCN_2(s)$ , according to



Write the Lewis formula of the cyanamide ion.

**M-4.** What are the major allotropes of carbon?

**M-5.** Given that graphite is converted to diamond under high pressures, what can you conclude about the relative densities of the two substances?

**M-6.** Explain on a molecular level why diamond is an extremely hard substance and graphite is slippery. Give applications of these properties.

**M-7.** Describe, using balanced chemical equations, how silicon is produced.

**M-8.** Describe the process of zone refining. What purity of silicon is obtainable by zone refining?

**M-9.** Use VSEPR theory (Chapter 8) to predict the shapes of the  $SiO_4^{4-}$  and  $SiO_6^{2-}$  ions.

**M-10.** Describe how photochromatic glass works.

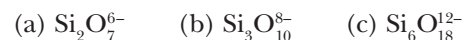
**M-11.** (a) Sketch a straight-chain silicate polyanion. How are the  $SiO_4^{4-}$  units linked to each other?

(b) Sketch the cyclic polysilicate ion,  $Si_6O_{18}^{12-}$ . In which mineral does this occur?

**M-12.** What are the four principal components of glass?

**M-13.** Describe the reaction by which  $HF(aq)$  etches glass.

**M-14.** Draw structures for the following ions:



**M-15.** Assuming a yield of 75%, how many metric tons of silicon can be obtained from 500 metric tons of silicon dioxide?