3.5: The Second Period of the Periodic Table

The first period of the periodic table is a short one consisting of only two elements, hydrogen and helium. Lithium, atomic number 3 begins the second period, which has 8 elements. Elements with atomic numbers 4-10, which complete this period, are discussed in this section.

Beryllium, Atomic Number 4

Like atoms of all the elements in the second period of the periodic table, **beryllium**, atomic number 4, atomic mass 9.012, has 2 inner shell electrons. Beryllium also has 2 outer shell electrons, so its Lewis symbol is

$$\mathbf{Be:}$$

In addition to 4 protons in their nuclei, beryllium atoms also have 5 neutrons. When the beryllium atom is oxidized to form a beryllium cation, the reaction is

$Be: \rightarrow Be^{2+} + 2e^-$ (lost to another atom) (3.5.1)

Since the beryllium atom needs to lose 2 electrons to reach the two-electron helium electron configuration, it produces doubly charged Be^{2+} cations.

Beryllium has some important uses in metallurgy. Melted together with other metals, a process that produces **alloys**, beryllium yields metal products that are hard and corrosion-resistant. Beryllium alloys can be blended that are good electrical conductors and that are nonsparking when struck, an important characteristic in applications around flammable vapors. Among the devices for which beryllium alloys are especially useful are various specialty springs, switches, and small electrical contacts. Beryllium has found widespread application in aircraft brake components where its very high melting temperature (about 1290˚C) and good heat absorption and conduction properties are very advantageous.

In a sense, beryllium is somewhat the opposite of a green element. This is because of its adverse health effects, including **berylliosis**, a disease marked by lung deterioration. Because of the extreme inhalation hazard of Be, allowable atmospheric levels are very low. Many workers were occupationally exposed to beryllium as part of the nuclear reactor and weapons industry in the U.S. in the decades following World War II. In recognition of the adverse health effects of occupational exposure to beryllium, in the late 1990s the U.S. Government agreed to compensate workers suffering occupational exposure to this metal.

Boron, a Metalloid

Boron, B, atomic number of 5, atomic mass 10.81, consists primarily of the isotope with 6 neutrons in addition to 5 protons in its nucleus; a less common isotope has 5 neutrons. Two of boron’s 5 electrons are in a helium core and 3 are outer electrons as denoted by

$$\dot{\mathbf{B}}\,\mathbf{:}$$

Boron is the first example of an element with properties intermediate between those of metals and nonmetals called **metalloids**. In addition to boron, the metalloids include silicon, germanium, arsenic, antimony, and tellurium, of which the most notable, silicon, is among the first 20 elements. In the elemental state, metalloids have a luster like metals, but they do not readily form simple cations. Unlike metals, which generally conduct electricity well, metalloids usually conduct electricity poorly, if at all, but can become conductors under certain conditions. Such materials are called **semiconductors** and are of crucial importance because they form the basis of the world’s vast semiconductor industry, which has given us small, powerful computers and a huge array of other electronic products.

Boron is a high-melting substance (2190˚C) that is alloyed with copper, aluminum, and steel metals to improve their properties. As a good absorber of neutrons, boron is used in nuclear reactor control rods and as a neutron-absorbing additive to the water that circulates through a reactor core as a heat transfer medium. Boron nitride, BN, is extraordinarily hard, as are some other compounds of boron. Boron oxide, B_2O_3, is an ingredient of heat-insulating fiber glass and boric acid, H_3BO_3, is used as a flame retardant in cellulose insulation in houses.

The Element of Life, Carbon

Carbon, C, atomic number 6, brings us to the middle of the second period of the periodic table. In addition to its 2 inner electrons, the neutral carbon atom has 4 outer electrons as shown by the Lewis symbol

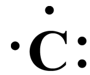


Figure 3.5 shows three organic compounds composed only of carbon and hydrogen (hydrocarbons), each containing 8 carbon atoms. These structures illustrate the bonding diversity of carbon.

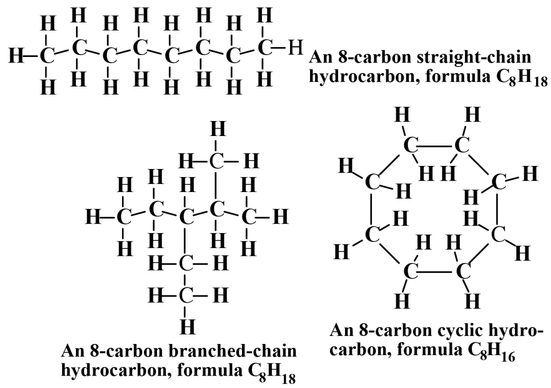


Figure 3.5. Illustration of the bonding versatility of carbon with three carbon-containing hydrocarbon compounds arranged as straight (continuous) chains, branched chains, and in a cyclic compound.

The ability of carbon atoms to bond with each other determines the properties of the several important and useful forms of elemental carbon. (Different forms of the same element are called allotropes.) Very fine carbon powder composes **carbon black**, which is used in tires, inks, and printer toner. Carbon atoms bonded in large flat molecules compose **graphite**, so soft and slick that it is used as a lubricant. Carbon treated with steam or carbon dioxide at elevated temperatures develops pores that give the carbon an enormous surface area. This product is **activated carbon** that is very useful in purifying foods, removing organic pollutants from water and removing pollutant vapors from air. Elemental carbon fibers are bonded together with epoxy resins to produce light composites so strong that they are used for aircraft construction. Bonded together in a different way that gives a very hard and rigid structure, carbon atoms produce diamond.

A particularly interesting class of carbon allotropes is that of **fullerenes** consisting of elemental carbon bonded in generally 5- and 6-membered rings to form spheres, ellipsoids, and tubes. The first of this class of elemental carbon discovered only in 1985 consists of aggregates of 60 carbon atoms bonded together in 5- and 6-membered rings that compose the surface of a sphere. This structure resembles the geodesic domes designed as building structures by Buckminster Fuller, a visionary designer. Therefore, the discoverers of this form of carbon named it buckminsterfullerene and the C60 balls, which resemble soccer balls in their structure, are commonly called “buckyballs.” Since the discovery of the C60 fullerene, many related forms have been synthesized, of which the most interesting may be very narrow carbon tubes called **carbon nanotubes**. (“Nano” is a prefix commonly assigned to materials in which the individual units have dimensions around 1 nanometer or 1 × 10-9 meters.) Carbon nanotubes have very interesting properties including some forms with an extraordinarily great length-to-diameter ratio of up to 132,000,000:1. Because of their unique dimensions, extraordinary strength, electrical properties, and ability to efficiently conduct heat, carbon nanotubes are of intense interest in materials science, nanotechnology, electronics, optics, and other high technology applications. However, special attention needs to be given to their potential toxicity.

Green Carbon from the Air

Carbon is present in the air as gaseous carbon dioxide, CO2. Although air is only about 0.04% CO2 by volume, it serves as the source of carbon for the growth of green plants. In so doing, the chlorophyll in plants captures solar energy in the form of visible light, represented *h*ν, and uses it to convert atmospheric carbon dioxide to high-energy glucose sugar, C6H12O6, as shown by the following reaction:

$$6\,CO_2 + 6\,H_2O \xrightarrow{h\nu} C_6H_{12}O_6 + 6\,O_2 \tag{3.5.1}$$

The carbon fixed in the form of C6H12O6 and related compounds provides the basis of the food chains that sustain all organisms. Organic carbon produced by photosynthesis in eons past also provided the raw material for the formation of petroleum, coal, and other fossil fuels. Now, as supplies of these scarce resources dwindle and as the environmental costs of their extraction, transport, and utilization mount, there is much renewed interest in photosynthetically produced carbon compounds as raw materials and even fuels. Despite the low levels of carbon dioxide in the atmosphere and the relatively low efficiency of photosynthesis, rapidly growing plants, such as some varieties of hybrid poplar trees, can produce enormous quantities of carbon compounds very rapidly and in a sustainable manner.

Nitrogen from the Air

Nitrogen, N, atomic number 7, atomic mass 14.01, composes 78% by volume of air in the form of diatomic N2 molecules. The nitrogen atom has 7 electrons, 2 contained in its inner shell and 5 in its outer shell. So its Lewis symbol is the following:

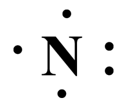


Figure 3.6.

Nitrogen gas does not burn and is generally chemically unreactive. Advantage is taken of the extreme chemical stability of nitrogen gas in applications where a nonreactive gas is needed to prevent fires and explosions. Although almost 80% of the air that people breathe consists of elemental nitrogen gas, people have died of asphyxiation by entering areas filled with nitrogen gas in which oxygen is absent. Since nitrogen gas has no odor, it does not warn of its presence.

Huge quantities of liquid nitrogen, which boils at a very cold -190˚C, are used in areas where cold temperatures are needed. This frigid liquid is employed to quick-freeze foods and for drying materials in freeze-drying processes, Biological materials, such as semen used in artificial breeding of animals, can be preserved in liquid nitrogen.

The atmosphere is an inexhaustible reservoir of nitrogen. However, it is very difficult to get nitrogen into the chemically combined form in which it occurs in simple inorganic compounds or proteins. This is because of the extreme stability of the N2 molecule, mentioned above. The large-scale chemical fixation of atmospheric nitrogen over a catalytic surface at high temperatures and pressure as represented by the reaction

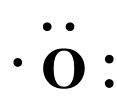
$$N_2 + 3\,H_2 \rightarrow 2\,NH_3 \tag{3.5.2}$$

was a major accomplishment of the chemical industry about a century ago. It enabled the large-scale production of relatively cheap nitrogen fertilizers that resulted in highly increased crop production, as well as the manufacture of enormous quantities of nitrogen-based explosives that made possible the unprecedented carnage of World War I. Despite the extreme conditions required for the preparation of nitrogen compounds by humans in the anthrosphere, humble bacteria accomplish the same thing under ambient conditions of temperature and pressure, converting N2 from the air into organically bound nitrogen in biomass. Prominent among the bacteria that do this are *Rhizobium* bacteria that grow symbiotically on the roots of legume plants, fixing atmospheric nitrogen that the plants need and drawing nutrients from the plants. Because of this ability, legumes, such as soybeans and clover grow well with less artificial nitrogen fertilizer than that required by other plants. One of the exciting possibilities with genetically modified plants is the potential to develop nitrogen-fixing varieties of corn, wheat, rice, and other crops that now lack the capability to fix nitrogen.

Nitrogen is an essential life element that is present in all proteins, hemoglobin, chlorophyll, enzymes, and other life molecules. It circulates through nature in the **nitrogen cycle** by which elemental nitrogen is incorporated from the atmosphere into biological material. Nitrogen-containing biomass is converted during biodegradation by bacteria to inorganic forms, which may be utilized as nutrient nitrogen by plants. Eventually, bacterial processes convert the nitrogen back to elemental N2, which is returned to the atmosphere to complete the cycle.

Oxygen, the Breath of Life

Oxygen, atomic number 8, atomic mass 16.00 is required by humans and many other living organisms. A diatomic nonmetal, elemental oxygen consists of O2 molecules and makes up 21% of the volume of air. Of its 8 electrons, the oxygen atom has 6 in the outer shell as represented by the Lewis formula:



Oxygen can certainly be classified as a green element for a number of reasons, not the least of which is that O in the atmosphere is there for the taking. Elemental oxygen is transferred from the atmosphere to the anthrosphere by liquifying air and distilling the liquid air, the same process that enables isolation of pure nitrogen. These two gases are also separated from air by adsorption onto solid surfaces under pressure followed by removal under vacuum. Pure oxygen has a number of applications including use as a gas to breathe by people with lung deficiencies, in chemical synthesis, and in oxyacetylene torches employed for welding and cutting metals.

Although the elemental oxygen molecule is rather stable, at altitudes of many kilometers in the stratosphere, it is broken down to oxygen atoms by the absorption of ultraviolet radiation from the sun as shown in Chapter 2, Reaction 2.13.1 As illustrated by Reaction 2.13.2, the oxygen atoms formed by the photochemical dissociation of O2 combine with O2 molecules to produce molecules of ozone, O3. The result is a layer of highly rarefied air containing some ozone over an altitude range of many kilometers located high in the stratosphere. There is not really much ozone in this layer. If it were pure ozone under the conditions of pressure and temperature that occur at ground level, the ozone layer would be only about 3 millimeters thick! This stratospheric ozone, sparse though it is, serves an essential function in protecting organisms on Earth's surface from the devastating effects of ultraviolet radiation from the sun. Were it not for stratospheric ozone, life as it is now known could not exist on Earth.

Ozone has a split personality as a green form of oxygen. As discussed above, ozone in the stratosphere is clearly beneficial and essential for life. But it is toxic to inhale at levels less than even one part per million by volume. Ozone is probably the most harmful constituent of air polluted by the formation of photochemical smog in the atmosphere at ground levels.

The most notable chemical characteristic of oxygen is its ability to combine with other materials in energy-yielding reactions. One such reaction with which most people are familiar is the burning of gasoline in an automobile,

$$2\,C_8H_{18} + 25\,O_2 \rightarrow 16\,CO_2 + 18\,H_2O + \text{energy} \tag{3.5.3}$$

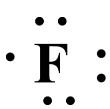
performed so efficiently that the combustion of only 1 gallon of gasoline can propel a full-size automobile more than 25 miles at highway speeds. Along with many other organisms, we use oxygen in our bodies to react with nutrient carbohydrates,

$$C_6H_{12}O_6(\text{glucose}) + 6\,O_2 \rightarrow 6\,CO_2 + 6\,H_2O + \text{energy} \tag{3.5.4}$$

to provide energy that we use. Whereas the combustion of a fuel such as gasoline occurs at red-hot temperatures, the "burning" of carbohydrates in our body occurs through the action of enzymes in the body at body temperature of only 37˚C.

The Most Nonmetallic Element, Fluorine

Fluorine, F, atomic number 9, atomic mass 19.00 has 7 outer electrons as shown by its Lewis symbol



Elemental fluorine consists of diatomic F2 molecules constituting a greenish-yellow gas. Fluorine is the most nonmetallic of all the elements. It reacts violently with metals, organic matter, and even glass! Elemental fluorine is a very corrosive poison that attacks flesh and forms wounds that heal very poorly. Because of its hazards, the practice of green chemistry seeks to minimize the generation or use of F2.

Fluorine is used in chemical synthesis. It was once widely employed to make Freons, chlorofluorocarbon compounds such as dichlorodifluoromethane, Cl2CF2, that were used as refrigerant fluids, spray can propellants, and plastic foam blowing agents. As discussed in Chapter 10, these compounds were found to be a threat to the vital stratospheric ozone layer mentioned in the discussion of oxygen above. They have now been replaced with fluorine-containing substitutes such as HFC-134a, CH2FCF3, which either do not contain the chlorine (Cl) that destroys stratospheric ozone or undergo destruction by atmospheric chemical processes near Earth's surface, and thus never reach the stratosphere.