Chapter 11
Patterns in the Chemical Elements

At the centre of chemistry is the study of the chemical elements and the compounds that they form. In this chapter, some of the trends and patterns among the elements will be discussed. Though there are many such similarities, it is important to realize each element has its unique properties. In this chapter, the focus will be on interesting and important main group elements.

11.1 Background

All matter consists of chemical elements and the compounds that they form. A question that many people ask is whether the chemical elements elsewhere in the universe are any different to those on Earth and whether there are many more chemical elements to be discovered.

In answer to the first question, as atoms of elements consist of combinations of protons, neutrons, and electrons, it would seem certain that the combinations are the same anywhere in the universe. The meteorites which impact the Earth, many of which are over four billion years old, contain minerals similar to those on Earth and consist of the same elements that we find on Earth.

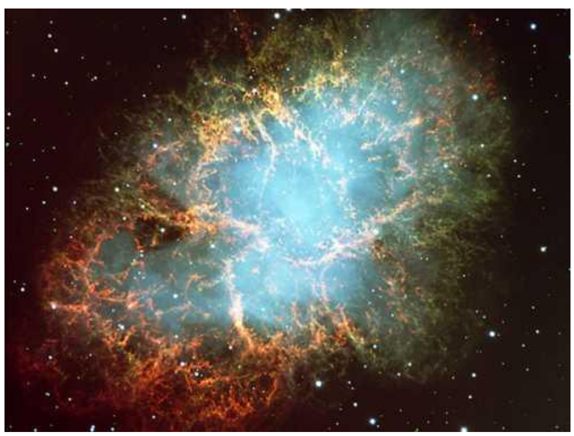

Figure 11.1 When the current universe began, only hydrogen and helium existed. It was through the explosions of stars, such as this one, that the nuclei fused together to form all the other chemical elements.

The answer to the second question is that it is certain that more chemical elements will be discovered. However, there is a problem. Atoms of the elements with high atomic number decompose over time, that is, the elements are radioactive. A simple explanation is that with increasing atomic number, more and more positively-charged protons are clustered in the nucleus. Positive particles repel each other and neutrons seem to act as a sort of ‘atomic glue’ holding them together in the nucleus. As the number of protons increase, so the number of neutrons has to increase at a faster rate in order to keep the nucleus together. For example, the most common isotope of carbon has six neutrons to accompany the six protons, while the most common isotope of lead requires 125 neutrons to stabilize the 82 protons. Beyond lead, it is impossible for any stable isotope of an element to exist (Figure 11.2).

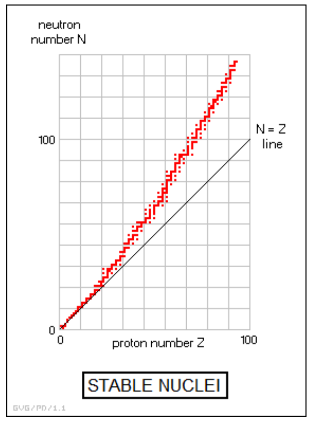

Figure 11.2 Stable nuclei

Uranium is the element with the highest atomic number that is found on Earth in measurable quantities. For example, half of all uranium atoms with 146 neutrons decompose in about four billion years. The more and more protons, the shorter and shorter will be the half-life, from hours, to seconds to milliseconds and even less. Thus even if a few atoms of new elements are synthesized, their existence will be so fleeting that their chemistry can never be studied in a traditional chemical laboratory. However, their synthesis will enable chemists to fill another box in the Periodic Table.

11.2 Group 1 (The Alkali Metals)

Metals are usually thought of as being hard, unreactive, and dense, with high melting points. The alkali metals are none of these: they are soft (Figure 11.3) and they have low densities and melting points. Densities increase down the Group while melting points decrease. For example, lithium has a density of 0.5 $g{\cdot}cm^{-3}$ (half that of water) and a melting point of 180ºC, while cesium has a density of 1.9 $g{\cdot}cm^{-3}$ and a melting point of only 29ºC. For comparison, iron – an 'average' metal – has a density of 7.9 $g{\cdot}cm^{-3}$ and a melting point of 1535ºC. Because the alkali metals are so different from 'typical' metals, alkali metals are sometimes called 'supermetals.'

Figure 11.3 Sodium, like the other alkali metals, is so soft that it can be cut by a knife.

A simple test for each of the alkali metals is to take a sample of an alkali metal compound and place it in a flame. Each of the alkali metals produces a characteristic colour:

Lithium – crimson red (Figure 11.4)
Sodium – yellow
Potassium – lilac
Rubidium – red-violet
Cesium – blue

Figure 11.4 The characteristic intense red of lithium in the flame test.

The alkali metals are very chemically reactive, always forming an ion of 1+ charge. In particular, they all react with the oxygen in the air. For example, lithium reacts with the oxygen in the air in a combination reaction to form lithium oxide:

$$4\ Li(s) + O_2(g) \rightarrow 2\ Li_2O(s)$$

Because the alkali metals react with the oxygen in the air, they are stored under oil.

In Chapter 9, Section 9.5, one of the sub-categories of single replacements was the reaction of the most reactive metals with water. Though lithium and sodium were the only alkali metals listed in Figure 9.9, in fact each one of the alkali metals react with water, the reaction becoming more violent from lithium through cesium. As an example, the chemical reaction of potassium with water is:

$$2\ K(s) + 2\ H_2O(l) \rightarrow 2\ KOH(aq) + H_2(g)$$

The reaction produces so much heat that the hydrogen gas then usually reacts with the oxygen in the air to form water in a combination reaction (Figure 11.5):

$$2\ H_2(g) + O_2(g) \rightarrow 2\ H_2O(l)$$

Figure 11.5 Potassium reacts so violently with water that the hydrogen produced catches fire and the flame is coloured lilac by traces of potassium ion.

SODIUM CHLORIDE

Seawater is a 3% solution of sodium chloride – together with many other ions in smaller proportions. Salt has been, and remains, a vital commodity. More sodium chloride is used by chemical industry than any other compound – in fact, about 150 million tonnes are consumed each year. Almost all sodium chloride is extracted from the many vast underground deposits, often hundreds of metres thick (Figure 11.6). These layers are the result of the evaporation of large lakes which evaporated hundreds of millions of years ago and which were then covered by other minerals.

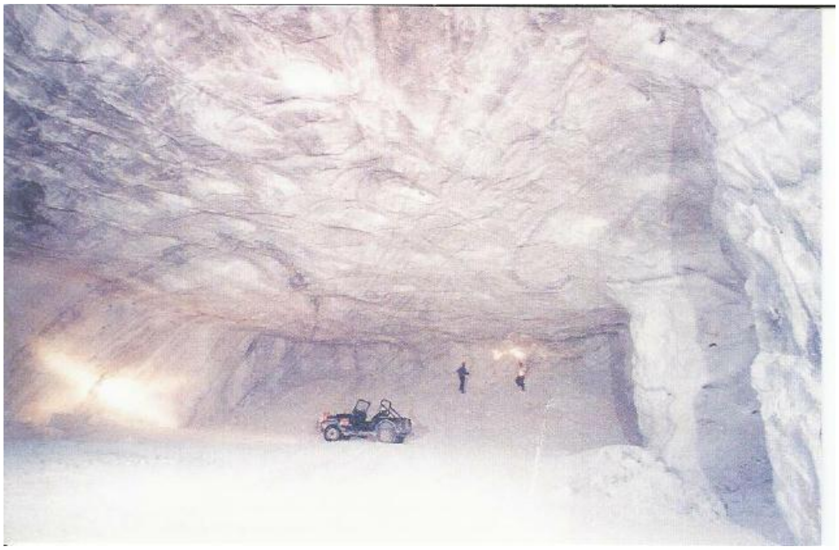

Figure 11.6 The interior of a salt mine

SODIUM CARBONATE AND SODIUM HYDROGEN CARBONATE

There are also underground deposits of sodium carbonate. About half of sodium carbonate is used in glass manufacture. The anhydrous compound, commonly called soda ash, is used to reduce the acidity of swimming pool water, while the decahydrate is commonly called washing soda.

One of the uses of sodium hydrogen carbonate is as a deodorant in refrigerators and carpets. It is also used in BC-type dry powder fire extinguishers. The sodium hydrogen carbonate smothers the fire and undergoes a decomposition reaction to give carbon dioxide gas and water vapour which assist in extinguishing the flames.

$$2\ NaHCO_3(s) \xrightarrow{\Delta} Na_2CO_3(s) + CO_2(g) + H_2O(g)$$

The main use of sodium hydrogen carbonate is in the food industry, to cause bakery products to rise. The sodium hydrogen carbonate is combined with calcium dihydrogen phosphate, $Ca(H_2PO_4)_2$. This mixture of ionic compounds, called baking powder, does not react in the solid phase, but when moistened, the compounds do react, one of the products being carbon dioxide gas, causing the bakery product to expand.

POTASSIUM CHLORIDE (POTASH)

Potassium chloride, mineral name, potash, is another alkali metal compound which is found in large underground deposits resulting from the evaporation of lakes millions of years ago. It is the potassium ion which is the crucial component, as potassium ion is needed for plant growth. As a result, potash is included in fertilizer mixtures. Canada holds the world's largest reserves of potash – mostly under Saskatchewan but also under New Brunswick.

BIOLOGICAL ASPECTS

We tend to forget that both sodium and potassium ions are essential to life. For example, we need at least 1 g of sodium ion per day in our diet. However, because of our addiction to salt on foods, the intake of many people is as much as five times that value. Excessive intake of potassium ion is rarely a problem. In fact, potassium deficiency is much more common; thus, it is important to ensure that we include in our diets potassium-rich foods such as bananas and coffee.

The alkali metal ions balance the negative charge associated with many of the protein units in the body. They also help to maintain the osmotic pressure within cells, preventing them from collapsing. Cells pump sodium ions out of the cytoplasm and pump potassium ions in. It is this difference in total alkali metal ion concentrations inside and outside cells that produces an electrical potential across the cell membrane.

The potential difference underlies many basic processes, such as the heart's generation of rhythmic electrical signals, the kidney's unceasing separation of vital and toxic solutes in the blood, and the eye's precise control of the lens's refractive index. Most of the 10 W of power produced by the human brain—awake or asleep—results from the Na^+/K^+–adenosine triphosphatase enzyme pumping potassium ion into and sodium ion out of brain cells. When we "go into shock" as a result of an accident, it is a massive leakage of the alkali metal ions through the cell walls that causes the phenomenon.

11.3 Group 2 (The Alkaline Earth Metals)

The alkaline earth metals are denser and harder than the alkali metals, though they are still comparatively chemically reactive. For example, heating magnesium powder or ribbon causes it to burn in the oxygen in the air in a combination reaction, also producing an intense white light (Figure 11.7):

$$2\ Mg(s) + O_2(g) \rightarrow 2\ MgO(s)$$

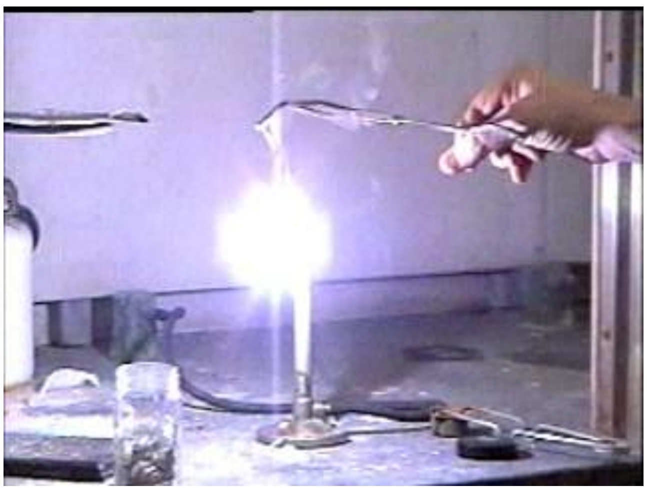

Figure 11.7 A piece of burning magnesium ribbon

All of the alkaline earth metals, except magnesium, react with water, though less violently than the alkali metals. For example, barium metal reacts with water in a single replacement reaction to give barium hydroxide solution and hydrogen gas:

$$Ba(s) + 2\ H_2O(l) \rightarrow Ba(OH)_2(aq) + H_2(g)$$

CALCIUM CARBONATE

Calcium is the fifth most abundant element on Earth. It is found largely as calcium carbonate in the massive deposits of chalk, limestone, and marble that occur worldwide. Chalk was formed in the seas, mainly during the Cretaceous period, about 135 million years ago, from the calcium carbonate skeletons of countless marine organisms. Limestone was formed in the same seas, but as a microcrystalline precipitate for calcium ions and carbonate ions present in seawater:

$$Ca^{2+}(aq) + CO_3^{2-}(aq) \rightarrow CaCO_3(s)$$

Some deposits of limestone became buried deep in the Earth's crust, where the combination of heat and pressure caused the limestone to melt. The molten calcium carbonate cooled again as it was pushed back up to the surface, eventually solidifying into the dense solid form that we call marble.

Figure 11.8 (left) The famous White Cliffs of Dover, England, consist of calcium carbonate in the form of chalk, while (right) the Taj Mahal in India was constructed from the harder limestone.

Calcium carbonate is a common dietary supplement prescribed to help maintain bone density. A major health concern today is the low calcium intake among teenagers. Low levels of calcium lead to larger pores in the bone structure, and these weaker structures mean easier bone fracturing and a higher chance of osteoporosis in later life.

Biological Aspects

The most important aspect of the biochemistry of magnesium is its role in photosynthesis. The magnesium ion, Mg^{2+}, is located in the middle of the chlorophyll molecule. It is the magnesium-containing chlorophyll, using energy from the Sun, which converts carbon dioxide and water into sugars and oxygen:

$$6\ CO_2(aq) + 6\ H_2O(l) \xrightarrow{\text{chlorophyll}} C_6H_{12}O_6(aq) + 6\ O_2(g)$$

Without the oxygen from the chlorophyll reaction, Earth would still be blanketed in a dense layer of carbon dioxide, and without the sugar energy source, it would have been difficult for life to progress from plants to herbivorous animals.

Both magnesium and calcium ions are present in body fluids. Mirroring the alkali metals, magnesium ions are concentrated within cells, whereas calcium ions are concentrated in the intracellular fluids. Calcium ions are important in blood clotting, and they are required to trigger the contraction of muscles, such as those that control the beating of the heart. In fact, certain types of muscle cramps can be prevented by increasing the intake of calcium ion.

11.4 Group 13

The only element of importance in this group is aluminum. Aluminum is the lowest-density comparatively unreactive metal. To extract aluminum from its ore, impure aluminum oxide, very large quantities of electrical energy are required. Thus the recycling of aluminum metal – such as aluminum cans – is a significant way of reducing energy consumption.

About half the density of iron and steel, aluminum is used for the bodies of most aircraft and for rapid-transit and railway vehicles. Aluminum is also a good conductor of heat, and for this reason it is used in cookware. However, copper is a much better heat conductor and so the better aluminum cookware has a copper-coated bottom.

BIOLOGICAL ASPECTS

Aluminum is the third most abundant element in the surface rocks, yet it is a highly toxic metal. Fortunately, in soils, aluminum ion forms insoluble compounds, such as aluminum hydroxide, $Al(OH)_3(s)$, minimizing its bioavailability. Much of the atmospheric pollution is acidic, and it is the hydrogen ions, $H^+(aq)$ from these acids which react with the solid aluminum hydroxide to give a solution of toxic aluminum ion in the ground water. The streams flow into lakes where the dissolved aluminum ion kills the fish.

$$Al(OH)_3(s) + 3\ H^+(aq) \rightarrow Al^{3+}(aq) + 3\ H_2O(l)$$

Human tolerance of aluminum is greater, but we should still be particularly cautious of aluminum intake. Part of our dietary intake comes from aluminum-containing antacids. Tea is high in aluminum ion, but the aluminum ions form inert compounds when milk or lemon is added. It is advisable not to inhale the spray from aluminum-containing antiperspirants because the metal ion is believed to be absorbed easily from the nasal passages directly into the bloodstream.

11.5 Group 14

Up to now, the elements discussed have been metals. Group 14 is the first to include the full range of element types: carbon, nonmetal; silicon and germanium, semimetals; and tin and lead, metals.

CARBON

In nature, there are two different forms of pure carbon: diamond and graphite, the two forms being the result of different arrangements of covalent bonds within the molecule. Such different forms of the same element are called ***allotropes***. Diamonds are formed naturally when the graphite allotrope is subjected to high pressures and temperatures deep within the Earth. Volcanic eruptions from millions of years ago brought some of these diamond deposits close to the Earth's surface in so-called "diamond pipes." It is now possible to undertake the same transformation in the laboratory but at present, only small industrial-use diamonds can be formed. For gemstone quality, diamond mines are the only source.

Diamond is the hardest naturally-occurring substance. This hardness results from each carbon atom being covalently-bonded to four neighbours and for this bonding to continue throughout each diamond crystal. Thus each diamond is one enormous molecule.

Graphite is a black slippery powdery or flaky substance. It is the only known nonmetallic substance to conduct electricity. In graphite, the carbon atoms are covalently bonded in layers. It is the sliding of these layers, one over another, that causes graphite to be so ‘slippery’ and hence a good solid lubricant. A mixture of graphite and clay is used in pencils and called ‘pencil lead’ (though there is no lead in it), the higher the proportion of graphite the softer the ‘lead’, while higher proportions of clay result in a harder ‘lead’. ‘HB’ is used to indicate the most common pencil lead mixture, while the ‘B’ range of numbers indicates softer, and the ‘H’ range, harder.

In 1982, scientists started wondering whether other forms of carbon might exist in interstellar space. Several teams of scientists performed experiments to try to synthesize other atomic arrangements of carbon atoms. Then in 1985, one such experiment produced carbon molecules containing spheres of 60 carbon atoms (see Figure 11.9). In addition to C_{60}, subsequent research showed that it was possible to make carbon spheres containing other numbers of carbon atoms, such as C_{70}. This family of spherical molecules has been named fullerenes. More recently, tubes of carbon atoms have been synthesized and these are known as carbon nanotubes.

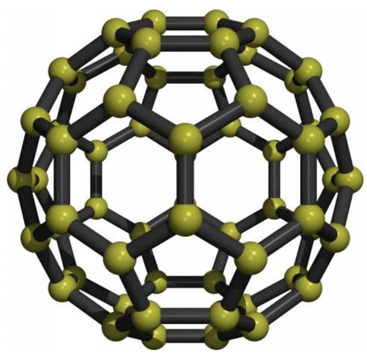

Figure 11.9 A ball-and-stick model of the C_{60} molecule

Though each allotrope has different physical properties, they have similarities in their chemical properties. In particular, each of the allotropes, when heated in air will burn to form carbon dioxide:

$$C(s) + O_2(g) \rightarrow CO_2(g)$$

Figure 11.10 The British chemist, Sir Humphry Davy, took the diamond from his wife's wedding ring and burned the diamond to prove that diamond was just a form of carbon.

SILICON DIOXIDE

On Earth, the large majority of rocks are silicates, that is, they contain silicate ions, such as SiO_4^{4-}, together with cations, such as aluminum and magnesium. Silicates are insoluble in water, and it is because of their insolubility that most of the mountain ranges – particularly the older ones – consist of silicate rocks. Also, silicon can be found as silicon dioxide, SiO_2, commonly given its mineralogical name of quartz (Figure 11.11). Silicon dioxide, like diamond, is very hard. The similarity can be explained in terms of a resemblance in chemical structure to diamond. Each quartz crystal, too, is a giant molecule, similar to the diamond structure, but with alternating silicon and oxygen atoms.

Figure 11.11 Crystals of silicon dioxide, commonly called quartz.

LEAD

This metal is soft, with a dull grey-black colour. It has a density about twice that of iron. Lead was used as water pipes because it was so easy to bend and, until recently was used in alloys as solder to hold water pipes together. The most important use now is as electrodes in batteries, such as those in motor vehicles. Lead is an inexpensive metal and the lead-acid battery can efficiently store electrical energy as chemical energy and then release it and subsequently be recharged.

11.12 A block of lead metal

Lead compounds were formerly used as pigments, for example, in paints. Lead(II) carbonate, $PbCO_3$, gives an intense white colour while lead(II) chromate, $PbCrO_4$, has an intense yellow colour and the latter is still often used for centre lines on highways (Figure 11.13). Until the twentieth century, lead(II) carbonate was used as a face powder for women to give a white tint to the skin.

Figure 11.13 Lead(II) chromate is an insoluble yellow-orange compound used as a paint pigment and for the middle-of-the-road markings on highways.

BIOLOGICAL ASPECTS

Carbon is the most crucial chemical element for life. The reason for this is that carbon atoms readily join together in chains. The ability of an element to covalently bond with itself is called ***catenation***. There are many biogeochemical cycles on this planet, the largest of which is the carbon cycle. Of the 2×10^{16} tonnes of carbon on and in this planet, most of it is "locked away" in the Earth's crust as carbonates, coal, and oil. Only about 2.5×10^{12} tonnes are available as carbon dioxide. Every year, about 15 percent of this total is absorbed by plants and algae in the process of photosynthesis, which uses energy from the Sun to synthesize complex molecules such as sucrose.

Silicon, as water-soluble silicates, is an essential element in our diet, though our daily needs are only about 10 mg per day. However, the dust of any silicate rock will cause lung damage called silicosis. As all silicates, except those of the alkali metals, are very insoluble in water (or body fluids), once the particles stick in the lungs, they are there for life. The irritation they cause produces scarring and immune responses that lead to the disease.

For many elements, the levels to which we are naturally exposed is many times smaller than toxic levels. For lead, however, there is a relatively small safety margin between unavoidable ingestion from our food, water, and air and the level at which toxic symptoms become apparent. About 95% of absorbed lead substitutes for calcium in our bones because lead(II) ions are about the same size as calcium ions. Lead interferes with the synthesis of hemoglobin; thus, it can indirectly cause anemia. At high concentrations, kidney failure, convulsions, brain damage, and then death ensue. There is also strong evidence of neurological effects, including reduced IQ in children exposed to more than minimal lead levels.

11.6 Group 15 (The Pnictogens)

For Group 15, nitrogen is the most important element. In addition, some of the properties of phosphorus will be described.

NITROGEN

There is only one allotrope of nitrogen, the diatomic molecule, N_2. Nitrogen is a colorless, odorless gas, that condenses to a liquid at −196ºC. The most important aspect of nitrogen chemistry is its lack of chemical reactivity. 78% of the Earth's atmosphere is nitrogen gas. In the context of the atmosphere, the lack of reactivity of nitrogen is very important, as nitrogen serves to reduce the combustibility of oxygen, the other major component in the atmosphere.

Nitrogen gas does not dissolve to a significant extent in water at normal pressures. However, the gas dissolves to a much greater extent under pressure. Thus the blood of deep-sea divers dissolves significant quantities of nitrogen. When the divers return to the surface, the gas bubbles out of their blood like the gas bubbles from a soft drink bottle when the cap is removed. These tiny bubbles – which commonly form around the joints – cause severe pain and can be fatal. To prevent 'the bends' as it is called, ascending divers have to return to the surface slowly. To avoid the problem, professional divers breathe a mixture of helium and oxygen gas instead.

PHOSPHORUS

There are two common allotropes of phosphorus: white and red (Figure 11.14). These two allotropes are very different from each other. White phosphorus, sometimes called yellow phosphorus, is a waxy-looking solid that spontaneously burns in air. To prevent combustion, white phosphorus is stored under water. Molecules of the white allotrope consist of four atoms in a pyramidal arrangement, P_4. Red phosphorus is a brick-red coloured powder which is stable in air. In red phosphorus, the atoms form long chains.

Figure 11.14 The two common allotropes of phosphorus: (left) white phosphorus, (right) red phosphorus.

BIOLOGICAL ASPECTS

Just as there is a carbon cycle, there is also a nitrogen cycle, for all plant life requires nitrogen for growth and survival. Between 10^8 and 10^9 tonnes of nitrogen is cycled between the atmosphere and the surface of the Earth each year. The dinitrogen in the atmosphere is converted by bacteria to compounds of nitrogen. Some of the bacteria exist free in the soil, but members of the most important group, the *Rhizobium,* form nodules on the roots of pea, bean, alder, and clover plants. This is a symbiotic relationship, with the bacteria providing the nitrogen compounds to the plants, and the plants providing a stream of nutrients to the bacteria. To do this at a high rate at normal soil temperatures, the bacteria use enzymes, biological catalysts.

Phosphorus is another element essential for life. For example, the free hydrogen phosphate and dihydrogen phosphate ions are involved in the blood buffering system. More important, phosphate is the linking unit in the sugar esters of DNA and RNA, and phosphate units make up part of ATP, the essential energy storage unit in living organisms. Finally, bone is a phosphate mineral, calcium hydroxide phosphate, $Ca_5(OH)(PO_4)_3$, commonly called apatite.

11.7 Group 16 (The Chalcogens)

For Group 16, oxygen is the most important element. In addition, some of the properties of sulfur will be described.

OXYGEN

There are two allotropes of oxygen, the diatomic molecule, O_2, and the triatomic molecule, O_3, commonly called ozone. (Di-)oxygen is a colorless, odorless gas, that condenses to a liquid at −183ºC. By contrast to nitrogen, the most important aspect of oxygen chemistry is its chemical reactivity. Oxygen does not burn itself, but many substances burn in oxygen. For example, oxygen is the only common gas that will ignite a glowing splint of wood.

21% of the Earth’s current atmosphere is oxygen gas. Early in the Earth’s history, there was no measurable proportion of oxygen in the atmosphere. As a result of photosynthesis, the proportion of oxygen in the atmosphere increased, reaching a maximum of about 35% about 100 million years ago. For about the last tens of million years, the proportion of oxygen in the atmosphere has been about 21%.

Like most gases, more oxygen dissolves in cold water than hot water. It is for this reason that the best fishing areas used to be in cold water areas such as the Labrador current, together with the Humboldt current off the coast of Chile and Peru.

Oxygen gas is released when a solution of hydrogen peroxide decomposes. When hydrogen peroxide solution is used to disinfect a wound, the hæmoglobin acts as a ***catalyst***, that is, the molecule speeds up the rate at which the reaction occurs. To indicate a catalyst, the name or formula is written over the reaction arrow:

$$2\ H_2O_2(aq) \xrightarrow{\text{hæmoglobin}} 2\ H_2O(l) + O_2(g)$$

Ozone, O_3, trioxygen, has very different properties to (di-)oxygen, O_2. Ozone is highly poisonous and it has a 'metallic' smell. At ground-level, ozone is a pollutant most often produced by the effect of sunlight on automobile exhausts. However, it is the layer of ozone in the upper atmosphere that protects us from the harmful high-energy ultra-violet rays.

SULFUR

Sulfur is a yellow solid. Its most common allotrope consists of rings of eight atoms, S_8 (though for chemical equations, the element is usually represented simply as 'S'). Sulfur is found naturally on the surface of the Earth around volcanic vents. What is most astounding, Jupiter's moon, Io, has vast quantities of sulfur, some as the liquid, spewing out in giant geysers with the plume rising up to 20 km into space.

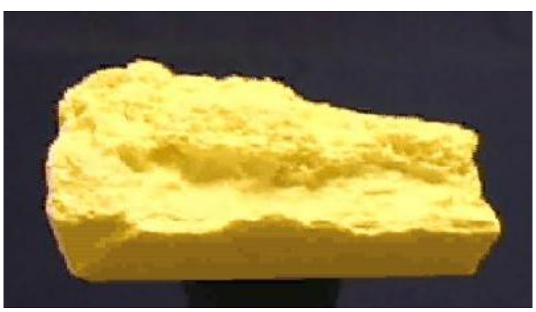
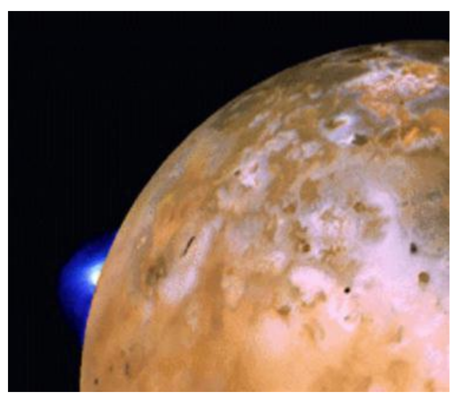

Figure 11.15 (left) a block of sulfur, (right) a molten sulfur geyser on the horizon of sulfur-covered Io, a moon of Jupiter.

BIOLOGICAL ASPECTS

Although humans can go for weeks without food and days without water, life ceases in a few minutes without oxygen. The average active adult inhales at least 10,000 L of air per day (about half the volume of a large gasoline truck). From this amount of air, about 500 L of oxygen are absorbed by the lungs. Because the body has no warning system for the lack of oxygen, one can lose consciousness because of a lack of oxygen without realizing what is happening.

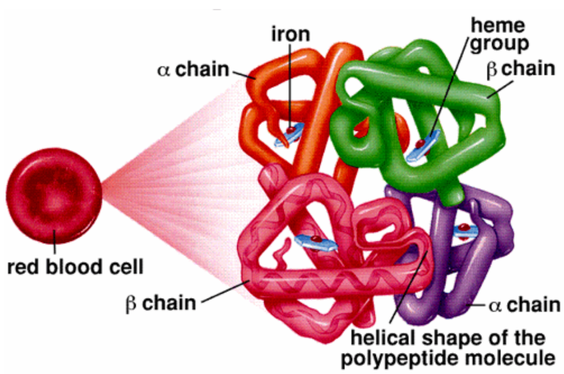

Figure 11.16 The hemoglobin molecule in red blood cells which pick up and transport oxygen molecules around an organism.

Sulfur is another essential element. Each molecule of the amino-acid cysteine contains one atom of sulfur (see Chapter 16, Section 16.8). Another sulfur-containing biological molecule is vitamin B_1 (thiamine). Many antibiotics, such as penicillin, cephalosporin, and sulfanilamide, are sulfur-containing substances. The majority of the simple sulfur-containing compounds have obnoxious odors, in fact, the odorous molecules from onion, garlic, and skunk all contain sulfur.

11.8 Group 17 (The Halogens)

For Group 17, each of the elements exists as a diatomic molecule: F_2, Cl_2, Br_2, and I_2. There are no other allotropes of any of these elements.

Figure 11.17 The common halogens (left) pale green chlorine gas, (centre) oily, red-brown bromine liquid, (right) black solid iodine with purple iodine vapour.

FLUORINE

Fluorine is a diatomic pale yellow-green gas which is so chemically reactive that it is sometimes called the *Tyrannosaurus Rex* of the chemical elements. In fact, it reacts with almost every other element. For example, it will react with a metal such as aluminum, to give solid aluminum fluoride:

$$2\ Al(s) + 3\ F_2(g) \rightarrow 2\ AlF_3(s)$$

and with a nonmetal, such as sulfur, to give gaseous sulfur hexafluoride:

$$S(s) + 3\ F_2(g) \rightarrow SF_6(g)$$

Thus whereas the reactivity of the alkali metals, and the alkaline earth metals, increased down their respective groups, for the halogens reactivity increases up the Group.

CHLORINE

This pale green diatomic gas is not quite as reactive as fluorine, but chlorine does react with most other elements. Like fluorine, chlorine is a very toxic gas, its toxicity leading to its use during the First World War as the first chemical warfare gas.

It is this same toxicity which has made it useful over the last hundred years as a means of purifying water. Bubbling small quantities of chlorine into domestic water supplies kills bacteria and viruses. Higher concentrations are used in swimming pool water as a disinfectant. In Quebec and most Francophone countries, ozone is used as the primary disinfectant. However, ozone decomposes over time to 'ordinary' oxygen, so the water in the water pipes would no longer remain sterile. :

$$2\ O_3(aq) \rightarrow 3\ O_2(aq)$$

For this reason, when ozone is used to purify water from reservoirs, some chlorine, $Cl_2(aq)$, still has to be added before it is sent out in the waterpipes.

A chemical test for chlorine is to add a solution of a soluble metal bromide. This is a single replacement reaction, discussed in Chapter 9, Section 9.5, involving the halogen replacement series. The solution will then go brown from the production of aqueous bromine. For example, aqueous chlorine will react with potassium bromide solution to give brown aqueous bromine and a solution of potassium chloride:

$$Cl_2(aq) + 2\ KBr(aq) \rightarrow Br_2(aq) + 2\ KCl(aq)$$

IODINE

Iodine, too, is diatomic, however it is a black metallic-looking solid which sublimes on warming to a violet-coloured gas. The gas is deposited as the solid on a cool surface to give beautiful crystals. Like all the other halogens, iodine is poisonous. The vapour can cause lung damage while dilute solutions are used as an antiseptic for wounds.

BIOLOGICAL ASPECTS

The chloride ion has a vital role in the ion balance in our bodies. It does not appear to play an active role but simply acts to balance the positive ions of sodium and potassium. However, covalently bonded chlorine is far less benign. Most of the toxic compounds with which we are currently concerned—for example, DDT and PCBs—are chlorine-containing molecules. The argument has been made to completely ban the production of chlorine-containing covalently bonded compounds. However, this would result in the elimination of many useful materials such as polyvinylchloride (PVC).

About 75 percent of the iodine in the human body is found in one location—the thyroid gland. Iodine is utilized in the synthesis of the hormones thyroxine and triiodothyronine. These hormones are essential for growth, for the regulation of neuromuscular functioning, and for the maintenance of male and female reproductive functions. Yet goiter, the disease resulting from a deficiency of the thyroid hormones (Figure 11.18), is found throughout the world, including a band across the northern United States, much of South America, and Southeast Asia; there are localized areas of deficiency in most other countries of the world. One common cause of the disease is a lack of iodide ion in the diet. To remedy the iodine deficiency, potassium iodide is added to common household salt (iodized salt).

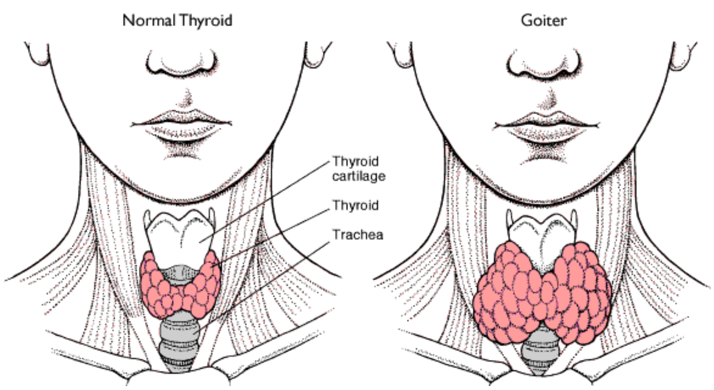

Figure 11.18 The normal thyroid and the swollen thyroid characteristic of goitre

11.9 Group 18 (The Noble Gases)

These gases are called the noble gases as it was believed that they did not form any chemical compounds. However, chemical compounds of krypton, xenon, and radon are now known. This group contains the gaseous element with the second-lowest density and the most dense gaseous element (radon). The density of a gas depends upon its molecular mass, so the pattern of densities of the noble gases can easily be explained. The density of helium is about 0.15 of that of air, while xenon is about 4.5 times that of air, and radon, 7.7 times that of air.

HELIUM

Helium is the next-lowest density gas to hydrogen. As helium is chemically unreactive, it is the gas used for party balloons and for balloons sent up to study the weather in the higher levels of the atmosphere.

The element name ending –*ium* is only used for metals – and helium. When astronomers were studying the light from the Sun about one hundred years ago, parts of the spectrum did not match any known substance on Earth. The element was named hel-ium in the belief that it was a metal. When the same element was subsequently discovered on Earth, it was found to be a nonmetal and a revised name of hel*ios* was proposed. Unfortunately, so many scientists were used to calling the gas 'helium' that the name never was corrected.

Helium also has the lowest condensation temperature of any gas, −269ºC. Liquid helium is used whenever very low temperatures are required, for example, to cool the magnets of MRI machines in hospitals. Helium is found naturally in a few underground deposits and there is some concern that the supplies of helium are becoming scarce.

XENON

Until the later part of the twentieth century, it was thought that no compounds of the noble gases could be prepared. However, a chemist at the University of British Columbia synthesized a chemical compound of xenon. Since then, more and more compounds of the noble gases have been produced. However, most of these have been compounds of xenon. Even now, no stable compounds of helium or neon have been produced.

BIOLOGICAL ASPECTS

None of the noble gases have any positive biological functions. Radon, however, is a significant health hazard. Radon isotopes are produced when radioactive uranium and thorium nuclei decompose. This process is happening continuously in the rocks and soils, and the radon produced normally escapes into the atmosphere. However, the radon formed beneath dwellings permeates through cracks in concrete floors and basement walls. It is not actually the radon that is the problem but the solid radioactive isotopes produced by its subsequent decay, such as polonium-218. These solid particles attach themselves to lung tissue, subsequently irradiating it with α-particles (helium nuclei) and β-particles (electrons), disrupting the cells and even initiating lung cancer. A properly ventilated house is always advisable not only to prevent possible radon accumulation but also more generally to flush out continuously all the air pollutants that are present in most modern well-sealed houses and offices.

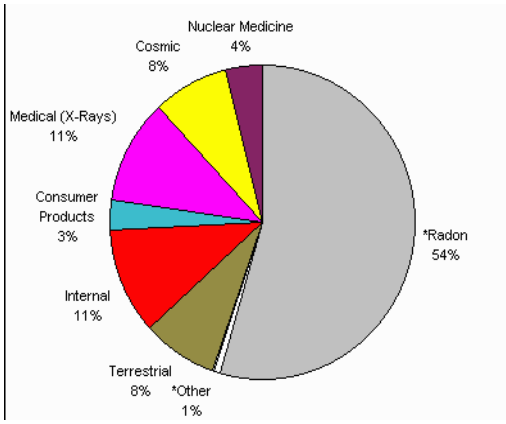

Figure 11.19 Sources of Radiation Exposure

11.10 Hydrogen

Hydrogen is the 'orphan' element. Some Periodic Tables have hydrogen placed at the top of Group 1, the alkali metals. This placing is because each hydrogen atom has one outer electron, as do atoms of the alkali metals. But hydrogen is not a highly reactive metal like the alkali metals! Other Periodic Tables place hydrogen at the top of Group 17, the halogens. The reason for this placement is that each hydrogen atom only needs one electron to complete its outer layer, as do atoms of each of the halogens. Though hydrogen is a diatomic gaseous nonmetal, as are the halogens, it is not highly corrosive, as are the halogens. Some Periodic Tables even have hydrogen at the top of both Groups 1 and 17. Because hydrogen is a unique element, it therefore makes more sense to place it on its own and discuss its chemistry separately.

Hydrogen is the least dense gas and it was used as the 'lifting gas' in airships. Unfortunately, hydrogen is a flammable gas, reacting with oxygen in a combination reaction to give water:

$$2\ H_2(g) + O_2(g) \rightarrow 2\ H_2O(l)$$

Hydrogen gas is formed in single replacement reactions of an acid with one of the mid-range of metals in the Activity series (discussed in Chapter 9, Section 9.5) from magnesium to lead. As an example, zinc metal reacts with hydrochloric acid to give zinc chloride and hydrogen gas:

$$Zn(s) + 2\ HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$$

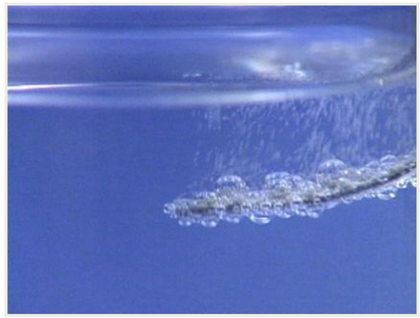
Figure 11.20 A strip of zinc under some hydrochloric acid with hydrogen gas bubbles forming on the surface of the strip.

11.11 Where Next?

In Chapter 12, we will focus on the behaviour of gases, including some of their chemical and physical properties, hence linking in with this chapter.