

TOPIC 12.

THE PERIODIC TABLE.

Reference web site: [http:// pearl1.lanl.gov/periodic/](http://pearl1.lanl.gov/periodic/)

Origin of the elements.

As discussed in Topic 1, there are 90 naturally occurring elements. In addition, there are about 20 other elements which have been produced synthetically but some of these exist only as very short-lived radioactive species which have been produced in extremely small quantities using high energy particle accelerators. How did the 90 naturally occurring elements originate? The most widely accepted version of the origin of all matter in the universe is the **BIG BANG THEORY** which proposes that in an instant, all of space, energy and matter which had been confined to a volume the size of a grain of sand at an infinitely high temperature underwent an explosive expansion. Within 10 minutes of the big bang, nuclei of mostly hydrogen and helium were formed from more basic particles and over the next million years these nuclei cooled enough to capture electrons and form atoms. After about a billion years, the gravitational attraction between atoms which were still mostly hydrogen and helium, lead to clumps of matter which by gravitational attraction gradually increased in size. With increasing size, the temperature of these clumps also increased and in some regions of space the larger clumps became hot enough to initiate fusion reactions between nuclei, forming stars which are in effect giant nuclear fusion reactors. Our own sun, like all stars, converts hydrogen to helium with the concurrent release of extremely large amounts of energy known as the **BINDING ENERGY** associated with the strong nuclear force which was discussed briefly in Topic 2. Within a few billion years vast numbers of new stars were formed and these in turn, through gravitational attraction, clustered to create galaxies, each of which contains enormous numbers of stars. By that point in time, the universe would have looked much as it does today. As the hydrogen fuelling the fusion reactions in a star is consumed, other fusion reactions can occur in which heavier elements form. Fusion reactions leading to new atoms of elements as heavy as iron all release energy and can continue to fuel a star. Ultimately, when a star has consumed most of its available fuel, it may simply cool and dim or in some cases it may initially implode and then undergo an enormous explosion which flings much of its constituent material and energy out into space in what is called a **SUPERNOVA EVENT**. It is only during supernovae that elements heavier than iron are formed. It is estimated that the present composition of the universe is 92.7% hydrogen atoms, 7.2% helium atoms and just 0.1% atoms of all the other elements. The shock waves of energy and material sent into space from supernovae may interact with existing clouds of gas, ice and dust to eventually form new stars and planets such as our solar system. Our sun was not one of the original stars in the universe but is probably a second or third generation star, formed in part from the energy and residues released by previous supernovae. Although already 5 billion years old, it still contains 71% hydrogen and 27% helium, so the sun will burn for several billion more years.

Thus on the basis of this theory, the elements which constitute all matter on earth were originally formed from stars that existed before our sun and which had consumed all their available hydrogen, converted it to helium and progressed to other nuclear fusion reactions that created heavier elements and finally underwent a supernova explosion in which more of the heavier elements were produced.

Isolation of the elements.

Reference web site: <http://homepage.mac.com/dtrapp/Elements/elements.html>

Most of the non-gaseous elements on earth are chemically combined with other elements as compounds. Few non-gaseous elements are found in the free state. For thousands of years gold, silver, copper, sulfur and carbon had been known because they do occur in the free form, although they were not necessarily recognised as elements - indeed the concept of an element as we know it today was not firmly established until the 18th century through the visionary work of Lavoisier. While some metals such as gold and silver are so unreactive that they can be found as free elements, most elements occur as compounds in **MINERALS**. Extraction of metals such as copper and tin from their ores by the process of **SMELTING** was probably accidentally discovered when minerals were used as fireplaces. The extraction process relied upon the use of bellows made from animal hides to increase the heat obtained from a fire to the point where decomposition could occur of rocks containing for example copper (malachite, a copper carbonate compound) and tin (cassiterite, an oxide of tin). Charcoal in the fire reduced the copper and tin compounds to the free elements. Later it was discovered that mixing about 9 parts of copper and 1 part of tin together and melting them produced an alloy called **BRONZE** which is much harder than either of its constituent elements, a feature exploited in the bronze age from about 3000 B.C. Later, iron was isolated from its ores by similar means - probably before 1000 B.C. Prior to 1600, the elements gold, silver, carbon, sulfur, copper, tin, lead, iron and mercury had been discovered by persons unknown. However arsenic, isolated by Albertus Magnus in 1250, is the first recorded instance of an attributed method of isolation of an element.

The prehistoric technical advance of using crude furnaces to smelt ores set the pattern for future discoveries of elements which resulted from newly devised processes. The rate of isolation of the elements was a process that tended to occur in steps where development of a new method or technology gave impetus to a flurry of discoveries, often followed by a quieter period prior to another new method or technology being devised. This is shown in the chart on page XII-6 which plots the number of known elements against dates of discovery based on the data given in the Tables on Pages XII-4 and 5. In the Table on Page XII-5, the dates of discovery of those elements known in prehistoric times are simply listed as <1600 while in the chart the prehistoric discoveries have been placed undated prior to 1600 along the axis.

Prior to 1600, the driver of isolation of elements such as copper, tin, lead and mercury was augmentation of the use of fire with bellows to create a furnace. Isolation of elements in the 18th century became increasingly rapid, boosted by several new technologies. Methods for analysing minerals reached a high state of development, particularly as a result of the advent of the blowpipe. This tool allows the ready decomposition of minerals on a carbon block, a prerequisite for analysis of their components. Elements discovered through the improved analytical techniques during the 18th century were cobalt, bismuth, platinum, zinc, nickel, manganese, molybdenum, tellurium, tungsten and chromium.

However, the greatest advance in new techniques during that period was the development of apparatus designed to handle gases. This not only allowed the isolation of the elements hydrogen, nitrogen, oxygen and chlorine as well as gaseous compounds such as carbon dioxide, but led to our current understanding of the nature of chemical processes and to fundamental laws such as the law of conservation of matter.

The development of the battery and its application to isolating elements via the method of electrolysis was exploited, particularly by Humphry Davy, in the early years of the 19th century. He used large batteries made from copper and zinc to electrolyse molten salts of elements from Groups I and II, isolating potassium, sodium, calcium and barium in rapid succession in 1807 - 1808.

Throughout the 19th century methods for analysing minerals continued to develop. These methods included the use of the blowpipe to obtain high temperatures in conjunction with acid digestion of minerals, selective precipitation of salts, gravimetric analysis and fractional crystallization allowing separation of pure salts from mixtures. Along with improved analytical techniques, the development of two new pieces of apparatus, the Bunsen burner and the spectroscope by Bunsen and Kirchoff respectively lead to discovery of elements initially by their atomic emission spectra. The Bunsen burner allowed high enough temperatures to produce atomic emission spectra from salts and the spectroscope allowed the unique pattern of spectral lines of each element to be observed and recorded. In 1860 they discovered the element caesium initially through its spectral lines which they found in samples extracted from minerals containing other Group I elements. In 1878 holmium and in 1879 samarium were first discovered by the presence of their spectral lines in extracts from ores. In 1868 Lockyer observed previously unknown spectral lines in sunlight and he attributed them to an unknown element which he called helium. Subsequently in the 1890's it was found that helium is present in our atmosphere along with the other noble gases neon, argon, krypton and xenon. Their discovery was in part prompted by Ramsay's observation that there was room for gaseous elements at the end of each Period of the Table. Removing all the known gases from air left a small, unreactive component which was called argon.

However, it required yet another technological advance before neon, krypton and xenon were isolated. This time the new technology was the ability to liquefy air. Fractional distillation of liquid air after gases such as oxygen, nitrogen, argon and carbon dioxide had been removed, left a small fraction which showed an unknown spectrum - this element was krypton. Careful fractional distillation of this remaining material revealed the presence of yet another gaseous element with an unknown spectrum - neon. Xenon was finally isolated by repeated fractionation of liquid krypton, again it was determined that they had obtained a new gaseous element by observing its spectrum. Thus a completely new family of hitherto unsuspected elements, the noble gases had been discovered as a result of combining spectroscopy and fractional distillation of liquid air.

The development of sensitive apparatus for measuring radioactivity by the Curies was central to their isolation of the elements polonium (1898) and radium (detected in 1898 but not isolated until 1903).

The chart shows a long gap until another rise in the rate of isolation of elements around 1940 when nuclear reactors were first invented and elements such as plutonium and americium were isolated for the first time. More recently, the development of high energy particle accelerators has been applied to production of synthetically produced elements, generally in microscopic quantities.

Isolation of the elements - chronological listing.

Prehistory	Au, Ag, C, S, Cu, Sn, Pb, Fe, Sb		
1250	As (but used earlier)	1866	F
1669	P	1875	Ga
1735	Co, Pt	1878	Yb, Ho
1746	Zn (but used earlier)	1879	Sm, Sc, Tm
1751	Ni	1880	Gd
1753	Bi (but used earlier)	1885	Pr
1755	Mg	1886	Dy, Ge
1766	H	1894	Ar
1772	O, N	1895	He
1774	Mn, Cl	1896	Eu
1778	Mo	1898	Po, Ra, Kr, Ne, Xe,
1781	W	1899	Ac
1782	Te	1900	Rn
1789	U, Zr, Be	1907	Lu
1790	Sr	1917	Pa
1791	Ti	1923	Hf
1794	Y	1925	Re
1797	Cr	1937	Tc
1801	Nb	1939	Fr
1802	Ta	1940	Np, Pu, At
1803	Os, Pd, Ce, Rh, Ir	1944	Cm
1807	Na, K	1945	Am
1808	Ba, B, Ca	1947	Pm
1811	I	1950	Bk, Cf
1817	Li, Se, Cd	1952	Es
1823	Al, Si	1953	Fm
1826	Br	1955	Md
1828	Th	1958	No
1830	V	1961	Lr
1839	La	1965	Rf
1843	Tb, Nd, Er	1970	Db
1844	Ru	1974	Sg
1860	Cs	1976	Bh
1861	Tl, Rb	1982	Mt
1863	In	1984	Hs

The Periodic Table

As each element has its own characteristic properties, this implies that one would need to be familiar with more than 100 different sets of chemical properties in order to understand the chemistry of all the elements. However the elements actually consist of families or Groups, each of which contains a number of elements that all share many similar properties. Thus by knowing the general properties of each family of elements, the task is made much easier. Further, recognition of the existence of these families has led to an understanding of why various properties are associated with each family. The arrangement of the elements as chemical families constitutes the **PERIODIC TABLE**, one of the fundamental cornerstones of chemistry which not only embodies the outward properties of elements, but also incorporates the inner atomic structure of their atoms.

Development of the Periodic Table.

The earliest suggested grouping of elements was simply on the basis of very obvious properties such as being shiny or malleable (classed as **METALS**) or not (classed as **NON-METALS**). Metals were further grouped as **COINAGE METALS** (silver, gold, copper) or as **REACTIVE METALS**. From this simple beginning, the modern Periodic Table evolved.

One of the first classifications into families was by Dobereiner (1829) who noted that there were often groups of three elements which shared similar properties, e.g.

Ca, Sr, Ba	reactive metals
Li, Na, K	very reactive soft metals
S, Se, Te	foul smelling hydrides
Cl, Br, I	highly corrosive non-metals
Fe, Co, Mn	hard metals, coloured salts

In each case, the atomic weight of the middle member was approximately the arithmetic mean of the other two.

As chemical knowledge increased along with the number of elements isolated, other bases for classification became possible. One basis tried was to arrange the elements in order of increasing atomic weight. Newlands (1864) observed that the chemical properties seemed to be repeated every 8 elements when this order was used, leading to his law of octaves: "*the span between repetitions in chemically similar species is an octave*". While Newlands's classification appeared valid for the lower atomic weight elements, at higher atomic weights there were many obvious absurdities with elements of very disparate properties being classed as a family. This problem arose because many elements had not yet been isolated in 1864, and no spaces had been left for them in his classification.

Dobereiner's suggestion of chemical families being related to atomic weight apparently inspired Mendeleef in 1868 to assemble a small fragment of a Periodic Table, based on more accurate atomic weights and chemical properties.

element	atomic weight	element	atomic weight	element	atomic weight
Cl	35.5	K	39	Ca	40
Br	80	Rb	85	Sr	88
I	127	Cs	133	Ba	137

In 1869 Mendeleef proposed a classification of the then known 65 elements which placed priority on allocating elements to each family on the basis of similar properties rather than atomic weight alone. He left blanks in families where discrepancies would otherwise appear and repositioned some elements, disregarding their accepted atomic weights and/or valencies in recognition that these could be in error. In 1871 he produced an updated version of the Table which included then unknown elements to which he gave names such as eka-aluminium (following aluminium: now gallium) and eka-silicon (following silicon: now germanium). Using this Table he predicted the properties of the missing elements based on the corresponding properties of the preceding and following element in the Group. His predictions were remarkably accurate when the missing elements were ultimately isolated.

The Periodic Law proposed by Mendeleef states: “*When arranged by atomic mass, the elements exhibit a periodic recurrence of similar properties*”.

The modern Periodic Table.

With the discovery of the sub-atomic particles and the subsequent knowledge of the structure of atoms, the fundamental basis for the periodic classification was realised to be *arrangement in order of increasing atomic number (the number of protons in the nucleus)*, rather than atomic weight (Moseley, 1913). The difference from atomic weight order is due to the various isotopes that contribute to the atomic weight (the weighted average of all naturally occurring isotopes of that element) so that an element which exists as an abundant heavy isotope but a lower atomic number would be out of order and appear in the wrong Group. There are only three instances where atomic weight order departs from atomic number order arising from the proportions of the different isotopes of each element. Atomic number is the number of protons in the atom’s nucleus and as atoms are electrically neutral, is numerically the same as the number of electrons in the atom. As the arrangement of electrons around the nucleus depends on how many electrons are present, then it is *the electronic structure of the atom which determines the properties of each element*. Thus what began as a classification solely on the basis of properties of elements was found to be a classification based on the atomic structure of the elements. Copies of the modern Periodic Table are given on the first and last pages of this book.

The Periodic Table in Review.

(i) The periods.

Each horizontal row or **PERIOD** of the Table starts with a Group 1 element and ends with a Group 18 element. [Note: There are several different Group numbering systems in use. One uses Roman numerals and extends from Group I to VII plus 0 or VIII for the noble gases while the current IUPAC system uses normal numbers and extends from 1 to 18. The IUPAC system is used in this Topic.] In between, there are various numbers of elements as follows:

1st period	H, He	2 elements
2nd period	Li → Ne	8 elements
3rd period	Na → Ar	8 elements
4th period	K → Kr	18 elements
5th period	Rb → Xe	18 elements
6th period	Cs → Rn	32 elements
7th period	Fr →	incomplete

Elements in each period *do not constitute a family* and there is no value in specifically committing to memory the members of each Period of the Table. Indeed, across each Period there is a steady change in properties from those elements classed as metals (on the left hand side) to those classed as non-metals (on the right hand side). Properties of metals compared with non-metals have been mentioned in previous Topics, but the following extended summary of physical and chemical properties of metals compared with non-metals is appropriate at this point.

<u>METALS</u>	<u>NON-METALS</u>
Good conductors of heat and electricity	Poor conductors
Malleable, ductile	Brittle, often powders or gases
Shiny appearance when freshly cut	Dull appearance
React with non-metals to form cations in salts	Form anions in salts
Dissolve in acids to form cations	Not soluble in acids
Ionic halides	Covalent halides
Have oxides which are ionic and dissolve in acids to form salts (basic oxides)	Have covalent oxides which are insoluble in acids but may dissolve in bases (acidic oxides).

The basis for the above properties and why they are associated with a metal or a non-metal are now well understood in terms of the structure of each element's atoms.

(ii) The Groups.

Each of the vertical columns of the Table headed Groups 1 through to 18 constitutes a chemical family of elements. Each family has many chemical properties in common. (You will recognise these as the eight groups that you committed to memory at the beginning of this course.) The chemical similarities within each Group are attributable to there being the same number of electrons in the outer level of the atoms of elements in that Group. For example, all atoms of Group 1 elements have 1 electron in their outer level; all Group 2 atoms have 2 electrons in the outer level and so on. This is the reason for there being common valencies within any given Group, a property to which Mendeleef gave priority in devising his version of the Table.

While the elements of each Group have many properties in common, normally there are also differences in properties within the Group. These differences typically are exhibited as a trend from the first to the last element within each Group. The most notable is that metallic properties (which are associated with the ease of removal of outer electrons) increase down the Group because the outer electrons are located at an increasing distance from the nucleus and thus attracted less strongly to it.

In addition, there is a large block of elements, called the d-block, located towards the middle of the table. You will also recognise some of these elements as those in the last section of Table 2 from Topic 1. Finally, there is another block of elements called the f-block located near the bottom of the Periodic Table.

The following notes provide an outline of the main chemical properties of each of the main Groups and also very briefly discusses the d-block and f-block elements.

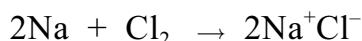
GROUP 1. Li, Na, K, Rb, Cs, (Fr) - known as the alkali metals because their hydroxides are all soluble in water.

They are all soft, very reactive metals which can be cut with a knife and tarnish rapidly so must be kept under oil to protect them from reacting with air.

They all react vigorously with water to form hydrogen gas, the reaction becoming increasingly violent down the Group. They form ionic compounds called "salts" with non-metals. For example, sodium and chlorine combine to form sodium chloride, a compound known as common salt or table salt. The elements in such compounds are no longer made up of electrically neutral atoms of free element, but instead are present as species called ions - atoms which have gained or lost electrons so as to have an electrical charge. When a salt forms, the Group 1 metal atoms have each had one electron removed, leaving them as +1 charged cations. As all Group 1 elements have a single outer electron, they all form only the M^+ ion in their compounds because this arrangement leaves the ions with the noble gas electron structure. In all compounds of Group 1 elements, each atom shows a combining power or valence of 1 only. To form cations with a 2+ or higher charge would require too much energy. The physical and chemical properties of Group 1 elements can be clearly related to this aspect of their atomic structure. The lone outer electron of Group 1 elements in the solid state can move from atom to atom easily when an electrical voltage is applied, causing these elements to be good conductors. As there is only the one outer electron available to bind each atom of the metal to its neighbours in the solid state, the metallic bond between the atoms is easily broken and causes the softness of Group 1 elements. In later groups where there are more outer electrons available to participate in metallic bonding, the elements become harder. The increased reactivity down the Group is a consequence of the outer electron being further from the nucleus and therefore more easily removed.

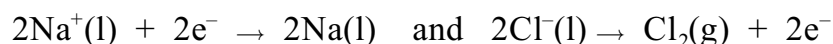
The atoms of a non-metal reacting with Group 1 elements gain the electrons that were removed from the metal atoms when their cations formed and so the non-metal atoms also become ions but with a negative charge (anions). The two oppositely charged ions are then held together by electrostatic attraction to form an ionic compound. These are characteristic reactions of all Group 1 elements.

For example, in the reaction of sodium with the non-metal chlorine, the process could be shown as



However, the charges on the Na^+ and Cl^- in such compounds are not usually shown in their formulas. It is a typical property of metals that they form ionic halides whereas non-metals form covalent halides.

When a salt is melted or when it is dissolved in water, it shows electrical conduction. This is because the ions are no longer held together in a solid crystal but instead, under the influence of an electrical voltage, move to the electrode of opposite charge and there undergo a reaction in which cations gain electrons and anions lose electrons to form free elements again. This process is called electrolysis and is the reverse of the reaction by which the original salt was formed. Electrolysis of molten sodium chloride will produce the elements sodium and chlorine at the electrodes according to the following equations.

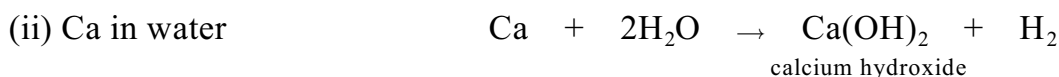


Reactions of this type, redox reactions, were examined in more detail in Topic 11. Sodium and potassium ions are important in the conduction of nerve impulses. Lithium ions provide a treatment for schizophrenia. Common compounds of Group 1 include sodium hydroxide (caustic soda), used in oven cleaners and as a starting material in many industrial processes and sodium hydrogencarbonate (NaHCO_3) which is used in cooking.

GROUP 2. Be, Mg, Ca, Sr, Ba, (Ra)

These are also metals, but are harder than Group 1 elements and their reactions are slower because more energy is required to remove the valence electrons. Beryllium is hard enough to scratch glass but at the bottom of the Group, barium is only slightly harder than lead. Increasing hardness compared with Group 1 elements can be attributed to the doubling of the number of outer electrons and a resulting increase in the strength of the metallic bonds. Apart from beryllium, they generally form salts with non-metals, always showing a valence of 2 in their compounds due to the presence of 2 outer electrons in all their atoms. Removal of the two outer electrons leaves all Group 2 elements with the noble gas structure. Beryllium forms covalent compounds with non-metals, a property more like that of a non-metal than a metal. The reason for this is that the Be atom is very small and the energy required to remove even one of its outer electrons is relatively high. Removal of both outer electrons from subsequent atoms in the Group becomes easier due to the increased size of the atoms and therefore less attraction between outer electrons and the nucleus and this explains the increased reactivity observed down the Group.

Some typical reactions of Group 2 elements -



Elemental beryllium is used as a component of specialised metals such as clockwork springs but compounds of beryllium are powerful carcinogens. Mg is the central atom in chlorophyll. The compound calcium carbonate (CaCO_3) is present in shells and calcium phosphate ($\text{Ca}_3(\text{PO}_4)_2$) is of particular importance as a component of teeth and bones. Large amounts of calcium carbonate (limestone) are used commercially in the manufacture of glass and cement. Radium has been used in radiotherapy treatment of cancer.

GROUP 13. B, Al, Ga, In, Tl

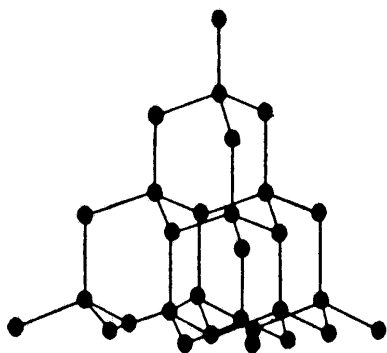
Boron is a non-metal, having a black powdery appearance and not having any of the usual properties of metals. Boron is a non-conductor of electricity and has a very high melting point (2040°C), indicating covalent network bonding of the boron atoms in the solid. It has an acidic oxide, typical of non-metals and it does not form ions when it reacts to produce compounds, but instead bonds by sharing electrons with the bonded atoms to form covalent compounds. This is the method by which non-metals bond to each other. The other elements of Group 13 mostly form ionic compounds. The Group 13 elements usually have a valence of 3 in their compounds, due to the presence of 3 outer electrons in their atoms.

Aluminium, the third most abundant element in the earth's crust, has considerable commercial application due to its very high strength to weight ratio and also its being a particularly good conductor of electricity. Aluminium is produced in large quantities by electrolysis of aluminium oxide ores such as bauxite. Although aluminium is very reactive, it forms an oxide layer (Al_2O_3) on its surface which protects the metal from further corrosion, making it useful as a building material. However, at high temperatures, aluminium burns vigorously.

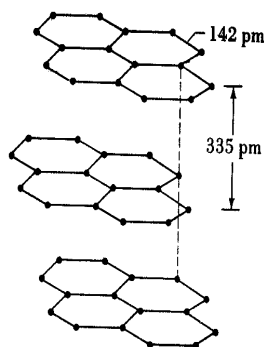
This Group shows clearly what is a general trend whereby the elements display increasing metallic properties down the Group. This trend is quite subtle in Group 1 where lithium does form a few covalent compounds and is more pronounced in Group 2 where beryllium forms compounds that are predominantly covalently bonded but retains many other properties of a metal including metallic bonding in the solid element. In Group 13, boron is essentially a non-metal in appearance and in physical and chemical properties. Aluminium forms both ionic and covalent compounds as do the other elements of this Group, but all exist as stable cations in solution. This trend from non-metallic to metallic properties down a Group continues to be apparent in Groups 14 - 16, attributable as discussed previously to the larger size of the lower atoms in each Group resulting in weaker attraction between nucleus and electrons and thus less energy being required to form cations.

GROUP 14. C (non-metal) Si, Ge (intermediate) Sn, Pb (metals)

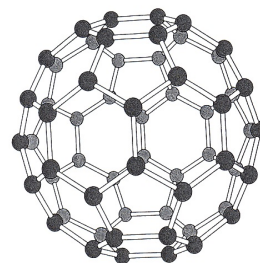
Carbon occurs as graphite and diamond as well as an amorphous (non-crystalline) form such as charcoal. Carbon also occurs as hollow, soccer-ball shaped arrangements of up to 76 or more C atoms known as fullerenes. These are examples of allotropes - different physical forms of the one element arising from different arrangements in the way their atoms are bonded.



The structure of diamond. Each large dot represents a carbon atom, and the lines represent bonds between atoms.



The structure of graphite.



The structure of a fullerene

Diamond contains a large number of carbon atoms joined by network covalent bonding. Each C atom is bonded to 4 other carbon atoms by covalent bonds which are pointing to the corners of a tetrahedron. Graphite contains carbon atoms which are bonded to 3 other carbon atoms, all in the same plane, and also has weak bonds to carbon atoms in the plane above and below. It is these weak bonds which cause graphite to easily peel off into flakes when it is used in "lead" pencils or in graphite lubricants. Graphite is exceptional among non-metals in being a conductor of electricity, also due to the mobility of the electrons that constitute the weak bonds between the planes of carbon atoms.

Carbon has an acidic oxide, CO_2 , forms only covalent bonds in compounds including its halides and apart from graphite, does not conduct electricity - typical properties of a non-metal. Carbon is the fundamental element in living cells, and is recycled through the carbon cycle. Many millions of carbon compounds exist as natural products or through laboratory or commercial synthesis. Because carbon has the ability to form molecules containing long chains of covalently bonded carbon atoms along with atoms of many other elements, almost unlimited numbers of carbon compounds are yet to be prepared. The chemistry of carbon compounds is called **ORGANIC CHEMISTRY**.

Important inorganic compounds of carbon include carbon dioxide, CO_2 , a product of **RESPIRATION** and the main **GREENHOUSE GAS** from burning fossil fuels. Carbon monoxide, CO , which results from the incomplete combustion of carbon compounds is a powerfully toxic substance. Fortunately, carbon monoxide rapidly converts to carbon dioxide in the atmosphere and does not accumulate to large concentrations.

Silicon is the fundamental element of mineral chemistry, and it makes up the majority of the earth's crust - usually in combination with oxygen as compounds called **SILICATES**. For example, the white sand on beaches is substantially silicon dioxide (silica), SiO_2 . About 87% of the earth's crust is made up of SiO_2 and related compounds. Silicon does not form such extensive chains through bonding as does carbon, and has nowhere near as many compounds. Apart from its role in mineral chemistry, elemental silicon is used to make solid state electronic components including photovoltaic cells for converting sunlight directly to electricity. Elemental silicon displays network covalent bonding in the solid state. Do not confuse the element silicon with the class of compounds called **SILICONES** which are polymers containing silicon bonded to hydrocarbon groups in particular. Silicon and germanium have some properties of both metals and non-metals so are often called **SEMI-METALS** or **METALLOIDS**.

While carbon and silicon do not form cations, germanium, tin and lead are more like metals in that they do form cations in compounds but they can also covalently bond. Group 14 again clearly illustrates this trend observed in most Periodic Table Groups towards an increase in metallic properties of the elements down each Group and both tin and lead in the solid state exhibit metallic bonding.

Tin and lead are both relatively low melting metals used together to form solder. Tin is used to plate the ubiquitous steel cans because, unlike iron, tin does not corrode readily.

Lead has anti-fungal properties which make it valuable as a component in paint, usually in the form of lead oxides. Concerns about health risks through ingestion of paint have caused lead to be deleted from most paints at present. Lead is also used in large amounts in the lead acid accumulator (i.e. the common rechargeable battery used in cars). There, alternate plates of lead metal and lead(IV) oxide, PbO_2 , are immersed in sulfuric acid, all contained in a plastic case. Electricity is provided by the redox reactions which take place at the electrodes. Redox reactions were discussed in detail in Topic 11.

Group 14 elements usually have valence = 4, but lead and tin in particular exist also in compounds with valence = 2, often as salts containing the ions Pb^{2+} or Sn^{2+} .

GROUP 15. N, P, (non-metals) As, Sb, (intermediate) Bi (metal)

Again, a transition from non-metal to metal is observed down this Group. Arsenic and antimony are usually regarded as metalloids. Nitrogen is one of the few elements to occur as a gas at room temperature and pressure. The other elements of Group 15 are solids. The elements nitrogen and phosphorus exhibit molecular covalent bonding, arsenic and antimony have network covalent bonding but bismuth has metallic bonding in the solid state as expected of a metal. Elements of this group commonly have valencies of 5 or 3 in their compounds. Compounds of nitrogen and phosphorus are of vital importance in biological systems - nitrogen as a component of all proteins and both nitrogen and phosphorus as components of DNA. Phosphate groups are an essential component of the system by which cells store and use energy. Elemental nitrogen is rather unreactive because it occurs as the highly stable N_2 molecule which constitutes 78% by volume of the atmosphere.

Simple compounds of nitrogen found in the natural environment as part of the nitrogen cycle include ammonia (NH_3), ammonium salts (containing the NH_4^+ ion), nitrate salts (containing the NO_3^- ion) and nitrite salts (NO_2^- ion). The process whereby atmospheric nitrogen is converted to such salts requires considerable energy input and is called **NITROGEN FIXATION**. It is accomplished by lightning strikes and by various microorganisms. Today, about 50 % of all nitrogen fixation is man-made through synthesis of ammonia from nitrogen and hydrogen gases.

Phosphorus occurs in four different allotropic forms, all of which react vigorously with oxygen in air, and thus must be stored under water. Phosphorus occurs most commonly in the natural environment as phosphates which contain the PO_4^{3-} ion.

Washing powders often contain phosphates.

Arsenic, antimony and bismuth occur less frequently and are of less importance in biological systems.

Being mostly non-metals, many compounds of Group 15 are covalently bonded, although compounds containing ions such as As^{3+} , Sb^{3+} , Bi^{3+} and N^{3-} commonly occur. Bismuth forms the usual salts containing the Bi^{3+} ion as expected of a metal.

GROUP 16. O, S, Se (non-metals) Te (intermediate) (Po - metal)

All except the rare and radioactive element, polonium, are non-metals or metalloids. They exist as 2- charged anions in compounds with metals and form covalently bonded compounds with other non-metals. Oxygen is the only member to occur as a gas, the others are solids at room conditions. Oxygen also occurs in small amounts in the atmosphere as the allotrope called **OZONE**. Ozone is dangerous to health when inhaled but plays a vital role in the upper atmosphere where it absorbs much of the harmful ultraviolet light that would otherwise impinge on the earth's surface. Tellurium has network covalent bonding in the solid state while elemental oxygen, sulfur and selenium are molecular covalently bonded. The common valencies are 6 and 2, but oxygen has the valence of 2 only.

Oxygen is one of the most reactive non-metals, forming oxides with most elements. The term **OXIDATION** originally referred to the reaction of substances with oxygen. Apart from occurring as the diatomic element O_2 (20.9 % by volume in air), oxygen also occurs extensively as part of numerous compounds in the earth's crust and in water. All aerobic organisms require elemental oxygen to survive as part of the process of respiration. Compounds of oxygen may be ionic, containing the oxide ion (O^{2-}), or covalent (e.g. H_2O , CO_2). Oxygen also occurs as the peroxide ion, O_2^{2-} , in ionic compounds or covalently bonded to non-metals such as in the antiseptic, hydrogen peroxide, H_2O_2 , which has the structural formula $\text{H}-\text{O}-\text{O}-\text{H}$. Sulfur occurs free in large deposits as the element and in compounds containing, for example, sulfide (S^{2-}) and sulfate (SO_4^{2-}) ions. Covalently bonded sulfur atoms are an important component of proteins, helping to maintain the required shape of enzymes. The compound hydrogen sulfide (rotten egg gas), H_2S , is a highly toxic gas which is generated by anaerobic bacteria. It is produced naturally by such bacteria in marshes and in sewage holding tanks where oxygen is excluded and can sometimes be detected in the exhaust fumes of modern motor cars.

Selenium and tellurium are very similar to sulfur in their properties and their hydrides, H_2Se and H_2Te , have an even more repulsive odour. Selenium is used commonly as a rectifier for converting AC to DC. Foods containing selenium compounds are in vogue at present because of the belief that they remove cancer-forming free radicals from the body.

All Group 16 elements form covalent compounds with other non-metals (e.g. $\text{H}-\text{O}-\text{H}$) and anions in ionic compounds with metals (e.g. $\text{Mg}^{2+}\text{O}^{2-}$).

GROUP 17 F, Cl, Br, I, (At) also known as the halogens.

Fluorine and chlorine are diatomic gases at room conditions, bromine is a liquid and iodine is a low melting solid. All exhibit molecular covalent bonding. The halogens are very reactive non-metals which can combine with metals to form ionic compounds containing the halide ion (F^- , Cl^- , Br^- or I^-). Species which can enter into such reactions are called **OXIDIZING AGENTS**, and all the halogens are therefore good **OXIDANTS**. Because of its oxidizing power, chlorine gas is introduced into the drinking water supply to oxidize organic contaminants such as bacteria and viruses which could be harmful, as well as to remove algae. Chlorine solutions can also be obtained for household use from the compound sodium hypochlorite ($NaOCl$) which is sold as bleach and also as an agent for maintaining clean swimming pool water. Iodine dissolved in alcohol (tincture of iodine) is used as a disinfectant for cuts of the skin, and is used to sterilise dairy equipment. Lack of iodide in food is one of the most common deficiency disorders in Australia where once it was mandatory to add iodide to table salt but this practice has ceased.

Some compounds of Group 17 elements are covalent (e.g. HCl) while others are ionic (e.g. Na^+Cl^-). They occur as anions in the latter case. The only ionic valence shown is 1 but, apart from fluorine, they have numerous other valencies when covalently bonded e.g. as the polyatomic anions in the salts $NaClO$, $NaClO_2$, $NaClO_3$, $NaClO_4$.

Important compounds of halogens include hydrogen chloride (HCl) which is a gas, but when dissolved in water it breaks down (ionizes) to form a solution of H^+ and Cl^- called hydrochloric acid. The halide ion Cl^- is an essential component of the nervous system and the fluoride ion, F^- plays an important role in developing strong enamel on teeth. As natural levels of F^- in drinking water are often too low, the salt sodium fluoride, NaF , is commonly added to drinking water to prevent tooth decay as a public health measure.

GROUP 18. He, Ne, Ar, Kr, Xe, Rn - the noble gases or inert gases.

All occur as monatomic gases at room conditions - the only monatomic elements. They have virtually no reactions due to the very stable atomic structure of 8 outer-shell electrons which all except helium possess and which is associated with very large energy requirements for the gain or loss of electrons required for reactions to occur.

Helium is found trapped in oil wells along with gas or crude oil. It occurs there because helium is a product of the radioactive decay of large, unstable nuclei such as uranium. Helium is an extremely useful substance because it has the lowest boiling point of any element. It constitutes a significant part of the mass of the sun where the nuclear fusion reaction of hydrogen to form helium provides most of the energy released.

The gases neon, argon and krypton are all present in the atmosphere in small amounts - argon to the extent of 0.9%. Neon and argon are used in situations where an inert atmosphere is required, such as in neon lighting tubes and in argon arc welding. The element radon is a powerful carcinogen which is ubiquitous in all minerals where it too arises from radioactive decay processes. Radon presents a particular hazard for miners involved in uranium mining, but is also released from the burning of coal and even from the clay of house bricks.

The "d-block" elements.

This is a block of elements located between Groups 2 and 13 following Ca, Sr and Ba. In the current IUPAC system of Group numbering, they are numbered as Groups 3 to 12. They first appear after Ca in the fourth period of the Table, and

there are ten d-block elements in each period where they occur. Note that there are no d-block elements in Periods 1, 2 or 3. They are all metals, most being hard with high melting points, although mercury is a liquid at room temperature.

Well known elements from this block include chromium, iron, manganese, nickel, cobalt, copper, zinc, cadmium, tungsten, platinum, silver, gold and mercury. Some of the special characteristics of these elements include the following:

- (i) They all form cations in salts (e.g. MnCl_2) but also can form covalent bonds to non-metals such as oxygen (e.g. Mn in the polyatomic ion MnO_4^-).
- (ii) They often have more than one ionic valence state (eg Fe occurs as the ions Fe^{2+} and Fe^{3+} in its compounds) as well as a number of valence states in covalent compounds.
- (iii) Their compounds are frequently coloured (eg copper(II) sulfate crystals - blue) whereas compounds of other metals are usually white.

The "f-block".

The f-block is located near the bottom of the Table and contains 14 elements in each of Periods 6 and 7. The first series, in Period 6, is called the lanthanides and the second, in Period 7, is called the actinides after the first element in each series. They are all metals and many of them have unstable nuclei. Particularly important are the elements uranium (U) and plutonium (Pu) because of their relevance to the nuclear fuel cycle.

Hydrogen

Hydrogen does not fit into any Group of the Table due to its unique property of having just one electron. Like Group 1 elements, it forms the +1 ion, but only in association with a molecule such as water because H^+ would be a naked proton, incapable of free existence. Hydrogen is also similar to Group 17 elements (halogens), existing as an H^- ion in ionic compounds, e.g. sodium hydride (NaH). However, ionic hydrides are unstable, reacting with water to form hydrogen gas and OH^- ions.

Metals vs non-metals - distribution throughout the table.

The metallic elements are located on the left hand section of the Table and the non-metals are on the right. Although the Periodic Table was originally devised purely on the basis of macroscopic properties, the true underlying basis for it is now understood in terms of the structure of the atoms of the elements. Metallic properties increase down each Group, even amongst those on the extreme left, because the atomic radius increases down each Group and outer electrons are more easily removed the further they are from the nucleus of the atom. The non-metals are on the right because they have more electrons in the same outer orbit which in turn are held more tightly by the increasing nuclear attraction from the greater number of protons, causing the atomic radius to decrease from left to right contrary to what one might have expected. Thus atoms of non-metals are more able to gain electrons to form anions rather than lose them to form cations. At the conclusion of each Period of the Table, the noble gases represent the most stable outer electron arrangement as the next element has its last electron located in a new energy level further from the nucleus and held less strongly, allowing that element to take on the properties of a metal again. This sequence of metal \rightarrow non-metal is repeated for each Period of the Table. The following table gives a guide to the general distribution of the metals compared with the non-metals in the Periodic Table. Between the metals and non-metals, towards the middle of the Table, some of the elements show properties of both and are called metalloids or semi-metals.

DISTRIBUTION OF METALS AND NON-METALS WITHIN THE PERIODIC TABLE

						H ○○ non-metal	Group 18
Group 1	Group 2	Group 13	Group 14	Group 15	Group 16	Group 17	He ○ non-metal
Li □ metal	Be □ metal	B ⊗⊗⊗ non-metal	C ⊗⊗⊗ non-metal	N ○○ non-metal	O ○○ non-metal	F ○○ non-metal	Ne ○ non-metal
Na □ metal	Mg □ metal	Al □ metal	Si ⊗⊗⊗ metalloid	P ○○ non-metal	S ○○ non-metal	Cl ○○ non-metal	Ar ○ non-metal
K □ metal	Ca □ metal	Ga □ metal	Ge ⊗⊗⊗ metalloid	As ⊗⊗⊗ metalloid	Se ○○ non-metal	Br ○○ non-metal	Kr ○ non-metal
Rb □ metal	Sr □ metal	In □ metal	Sn □ metal	Sb ⊗⊗⊗ metalloid	Te ⊗⊗⊗ metalloid	I ○○ non-metal	Xe ○ non-metal
Cs □ metal	Ba □ metal	Tl □ metal	Pb □ metal	Bi □ metal	Po □ metal		Rn ○ non-metal
Fr □ metal	Ra □ metal						

Bond type in the element

□ metallic

⊗⊗⊗ network covalent

○○ molecular covalent

○ monatomic (noble gases only)

Objectives of this Topic.

When you have completed this Topic, including the tutorial questions, you should have achieved the following goals:

1. Understand that the Periodic Table was originally devised on the basis of families of elements with similar properties, arranged in order of increasing atomic weight (but subsequently found to be increasing atomic number).
2. Know the electronic basis for the similarities and trends within Groups of the Period of the Table.
3. Commit to memory the elements of Groups 1, 2, 13-18.
4. Know the general distribution of metals and non-metals within the Table.
5. Have some familiarity with the main chemical properties of each Group.

SUMMARY

This classification of elements into groups with similar chemical and physical properties was begun prior to any knowledge of the structure of atoms and when the existence of undiscovered elements was overlooked. Purely on these bases, it was apparent that some elements having similar properties could be grouped together as families. With the measurement of atomic weights of elements came attempts to find correlations of properties with atomic weight order, but these were hampered by the many elements that had not been isolated at that time and also by inaccurate atomic weight determinations. Mendeleef recognised these deficiencies and turned them to advantage, still using atomic weight order but giving primacy to assigning elements with similar properties to the same Group. He left gaps in the arrangement of the elements where needed and predicted which elements were still to be discovered and their likely properties. This arrangement of the elements, called the Periodic Table, was further refined when the structure of the atom was elucidated and it was then realised that the order of the elements in the Periodic Table should be atomic number order rather than atomic weight order. From today's knowledge of atomic structure, the reason for similarities of chemical properties within any Group of elements has been clearly established as a consequence of each element in any Group having the same outer-shell electron arrangement.

However, gradations in properties within a Group are normally observed, with the most metallic elements being at the bottom. Within any row (Period) of the Table, the elements on the left are metals and on the right are non-metals, a gradual increase in non-metallic properties from left to right being exhibited. Some elements towards the middle of the Table show properties of both metals and non-metals. These trends can also be related to aspects of atomic structure, in particular the increasing size of atoms down each Group which leads to less energy being required to remove electrons and thus form cations - a property of metals. Elements to the right of the Table are non-metals because they have more electrons contained within the same outer level of the atom and are accompanied by a corresponding increase in the number of protons in the nucleus, resulting in all the electrons being held more tightly. Thus more energy is required to remove electrons from non-metals - instead, formation of anions or covalent bonding are their energetically preferred options.

The Periodic Table is an elegant example of how observations and the collection of data, development of hypotheses, predictions made from them and testing of each hypothesis followed by discarding or refining it leads to increased understanding in science.

Before commencing the questions associated with Topic 12, complete any remaining questions from Topics 11, 7, 8, 9 and 10 in that order.

TUTORIAL QUESTIONS - TOPIC 12

1. List the physical properties of metals compared with non-metals.
2. How do the chemical properties differ for metals compared with non-metals with regard to the following:
 - (a) Type of ion formed in salts
 - (b) Reaction with acids
3. Disregarding hydrogen and helium, for the Periodic Table Groups 1, 2, 13 - 18, what electronic structural feature do all the atoms in any given Group have in common?
4.
 - (a) Write the names and symbols for the elements of Group 1
 - (b) List some properties of Group 1 elements which indicate they are all metals.
 - (c) What valence do all Group 1 elements exhibit in their compounds?
5.
 - (a) Write the names and symbols of the elements of Group 2.
 - (b) What properties would indicate that beryllium could in some respects be more appropriately regarded as a non-metal?
 - (c) Aside from beryllium, list properties of the other members of Group 2 which indicate they are metals.
 - (d) What valence do all Group 2 elements show in their compounds?
6.
 - (a) Write the names and symbols of the elements of Group 13.
 - (b) Is boron a metal or a non-metal? Give reasons for your answer.
 - (c) Why is aluminium a useful structural material even though it reacts readily with water and oxygen?
 - (d) What is the usual valence shown by elements of Group 13 in compounds?
7.
 - (a) Write the names and symbols of the elements of Group 14.
 - (b) Which elements of Group 14 could be best regarded as metals rather than non-metals? Give evidence to support your answer.
 - (c) Diamond and graphite are both forms of pure carbon. Compare their physical properties and explain the differences in terms of arrangement of their atoms.

8.
 - (a) Write the names and symbols of the elements of Group 15.
 - (b) What type of bonding is present in elemental nitrogen?
 - (c) Give the names and formulas for some species to which nitrogen is converted in the process of nitrogen fixation.
 - (d) Summarise the metallic/non-metallic properties of the Group 15 elements.
 - (e) In what way does bismuth differ from the other members of Group 15?

9.
 - (a) Write the names and symbols for the elements of Group 16.
 - (b) Summarise the metallic/non-metallic properties of Group 16. List properties which support your answer.

10.
 - (a) Write the names and symbols for Group 17 elements (halogens).
 - (b) Explain why the halogens all have low melting and boiling points.
 - (c) Why are halogens strong oxidising agents?
 - (d) List all the elements that occur naturally as diatomic molecules.
 - (e) What properties confirm that Group 17 elements are all non-metals?

11.
 - (a) Write the names and symbols for Group 18 elements (noble gases).
 - (b) Why do all the elements of Group 18, the noble gases, occur as monatomic gases in nature?

12.
 - (a) Where in the Table do the elements of the d-block appear?
 - (b) While the d-block elements are all metals, these elements show some properties in general that differ from the metals of Groups 1 and 2. What are these different properties?

13. Summarise the occurrence of metals and non-metals in the Periodic Table.

14. Where in the Periodic Table are those elements which are (a) most easily oxidised and (b) those most easily reduced located? Explain why this is so.

ANSWERS TO TUTORIAL TOPIC 12

1. Metals: shiny when freshly cut, malleable, ductile, good conductors of heat and electricity

Non-metals: dull solids, powders or gases, brittle, poor conductors of heat and electricity
2. (a) Metals form cations in reactions that produce salts while non-metals form anions in those reactions.

(b) Many metals react with acids forming cations as part of a salt while non-metals do not react with acids.
3. All the elements in any Periodic Table Group have the same arrangement of electrons in their outer level.
4. (a) lithium (Li), sodium (Na), potassium (K), rubidium (Rb), caesium (Cs)

(b) They all have the physical properties listed for metals above; they all form cations in reactions with acids to produce salts.

(c) Group 1 elements always have a valency of 1 in their compounds.
5. (a) beryllium (Be) magnesium (Mg) calcium (Ca) strontium (Sr) barium (Ba)

(b) Beryllium does not form cations in its compounds, instead it is usually covalently bonded.

(c) They all display the physical properties of metals given above; they all form cations in reactions with acids to produce salts.

(d) Group 2 elements always have a valency of 2 in their compounds.
6. (a) boron (B), aluminium (Al), gallium (Ga), indium (In), thallium (Tl)

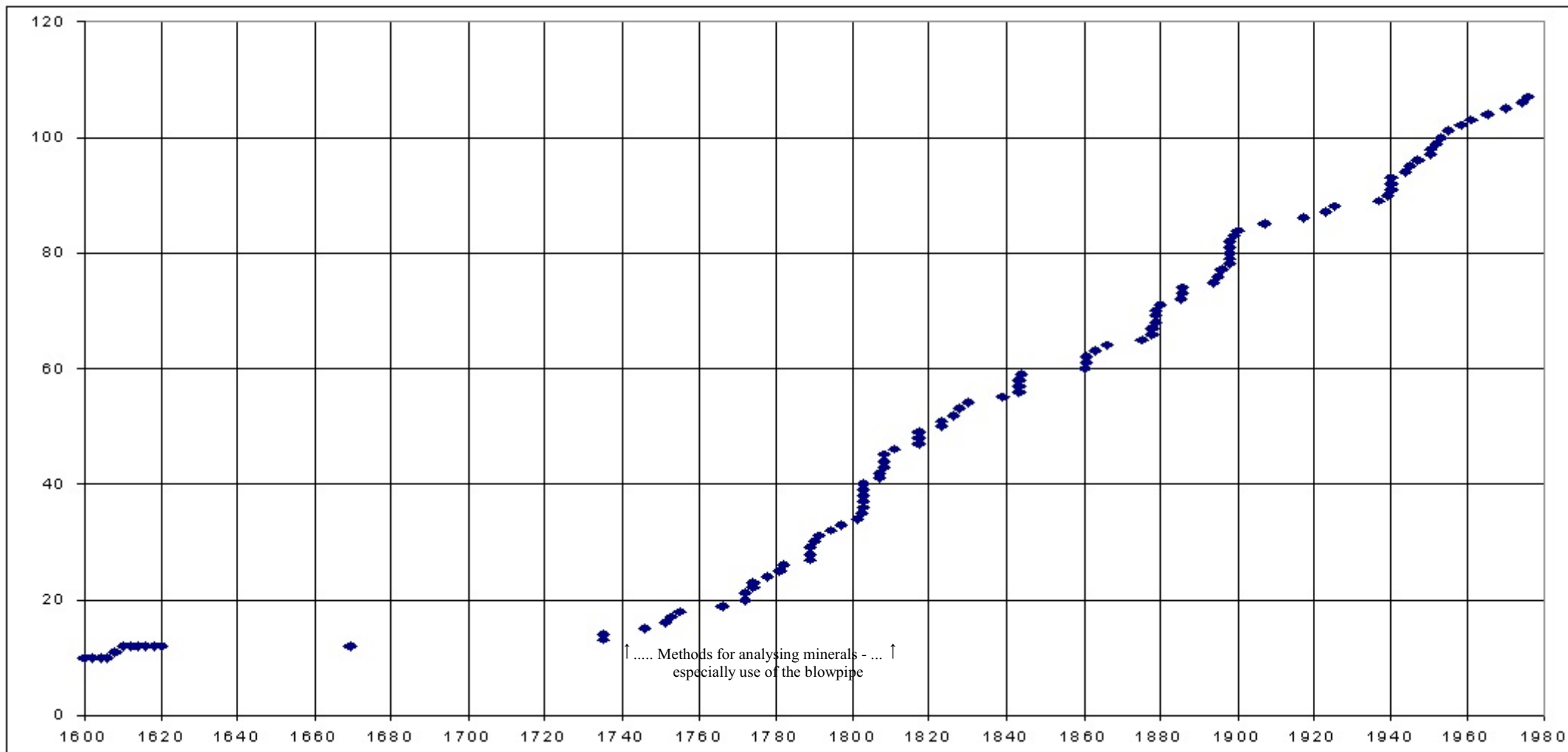
(b) Boron is a non-metal as it is a black powder, does not form salts with acids and is only covalently bonded in its compounds.

(c) Aluminium reacts with oxygen in the air to produce an oxide which adheres strongly to the surface of the metal and protects it from further corrosion in the atmosphere.

(d) Group 13 elements mostly have a valency of 3 in their compounds.

7. (a) carbon (C), silicon (Si), germanium (Ge), tin (Sn), lead (Pb)
- (b) tin and lead behave more like metals than the rest of Group 14 because they have the physical properties of metals and form cations in some of their compounds with non-metals.
- (c) Carbon atoms in diamond are arranged in a very stable tetrahedral structure with each C atom bonded to four other C atoms. The very stable structure of diamond imparts the properties of considerable hardness and non-conduction of electricity. In graphite, each C atom is bonded to just three other C atoms in a planar arrangement, leaving one unused valence electron on each atom. These unused electrons form weak partial bonds to C atoms in the planes above and below. Consequently graphite is soft and the planes of C atoms are easily peeled apart (as in lead pencils). Also graphite can conduct electricity because the electrons between the planes are so weakly held that they are mobile under the influence of an electrical voltage.
8. (a) nitrogen (N), phosphorus (P), arsenic (As), antimony (Sb), bismuth (Bi)
- (b) Elemental nitrogen consists of molecules consisting of two N atoms covalently bonded by a triple bond. This is termed molecular covalent bonding.
- (c) ammonia (NH_3), ammonium ion (NH_4^+), nitrogen dioxide (NO_2), nitrate ion (NO_3^-), nitrite ion (NO_2^-)
- (d) Nitrogen and phosphorus are clearly non-metals. Arsenic and antimony have properties of both metals and non-metals. Bismuth is more metallic than non-metallic.
- (e) Bismuth forms many ionic compounds containing the Bi^{3+} cation and it has metallic bonding in the elemental state.
9. (a) oxygen (O), sulfur (S), selenium (Se), tellurium (Te)
- (b) All the elements of Group 16 are non-metals. They have the usual physical properties of non-metals, do not dissolve in acids and they form anions when in compounds with metals. Their compounds with other non-metals are covalently bonded.
10. (a) fluorine (F), chlorine (Cl), bromine (Br), iodine (I)
- (b) The halogens occur as diatomic molecules which only have weak forces of attraction between their molecules. Consequently it requires relatively little energy (thus lower temperature) for the molecules to separate into the liquid phase from the solid (melt) or from the liquid phase into the gas (boil).

- (c) The atoms of halogens only need to gain one extra electron to become isoelectronic with a noble gas. Species which readily gain electrons are good oxidants.
- (d) H_2 , N_2 , O_2 , F_2 , Cl_2 , Br_2 , I_2
- (e) Apart from their physical properties, the halogens all form anions in salts and covalent compounds with other non-metals.
11. (a) helium (He), neon (Ne), argon (Ar), krypton (Kr), xenon (Xe), radon (Rn)
- (b) The outer electron level of the atoms of all the noble gases is filled with 8 electrons (or 2 electrons for helium), the arrangement that gives the noble gases their stability. To form diatomic molecules for example, more electrons would have to enter this level in forming covalent bonds between the atoms and this is not possible as the level is already filled.
12. (a) The d-block elements are located in the Periodic Table between Group 2 and Group 13. Where they occur, there are 10 d-block elements in each Period.
- (b) Among the properties that distinguish d-block elements from the metals of Groups 1 and 2 are that they frequently show a range of valencies and often have coloured compounds.
13. Metals are located from the left hand side starting with Group 1 and extending to include the d-block. The non-metals are located on the right hand side of the table extending to boron in Group 13 and carbon in Group 14. Metallic properties increase down each Group of the Table.
14. (a) The strongest oxidants are located in the top, right hand corner of the Periodic Table, excluding the noble gases. Fluorine is the strongest oxidant followed by oxygen and chlorine. Elements in this region of the table are only one or two electrons short of having the stable structure of the nearest noble gas, so energy is released when the atom captures an electron in the process of oxidizing another atom.
- (b) The strongest reductants are located in the bottom, left hand part of the Table where the outer electrons of elements are further out from the nucleus and not held so tightly by attraction to it. Consequently relatively small amounts of energy are required for an oxidant to remove an electron. Group I (1) elements are the best reductants as they only need to lose one electron to become isoelectronic with the nearest noble gas.



DISCOVERY OF THE ELEMENTS - NUMBER OF KNOWN ELEMENTS vs YEAR
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DATES OF DISCOVERY OF THE ELEMENTS

											1 H 1766							2 He 1895
3 Li 1817	4 Be 1798											5 B 1808	6 C <1600	7 N 1772	8 O 1772	9 F 1866	10 Ne 1898	
11 Na 1807	12 Mg 1755											13 Al 1823	14 Si 1823	15 P 1669	16 S <1600	17 Cl 1774	18 Ar 1894	
19 K 1807	20 Ca 1808	21 Sc 1879	22 Ti 1791	23 V 1830	24 Cr 1797	25 Mn 1774	26 Fe <1600	27 Co 1735	28 Ni 1751	29 Cu <1600	30 Zn 1746	31 Ga 1875	32 Ge 1886	33 As 1250	34 Se 1817	35 Br 1826	36 Kr 1898	
37 Rb 1861	38 Sr 1790	39 Y 1794	40 Zr 1789	41 Nb 1801	42 Mo 1778	43 Tc 1937	44 Ru 1844	45 Rh 1803	46 Pd 1803	47 Ag <1600	48 Cd 1817	49 In 1863	50 Sn <1600	51 Sb <1600	52 Te 1782	53 I 1811	54 Xe 1898	
55 Cs 1860	56 Ba 1808	57-71	72 Hf 1923	73 Ta 1802	74 W 1781	75 Re 1925	76 Os 1803	77 Ir 1803	78 Pt 1735	79 Au <1600	80 Hg <1600	81 Tl 1861	82 Pb <1600	83 Bi 1753	84 Po 1898	85 At 1940	86 Rn 1900	
87 Fr 1939	88 Ra 1898	89-103	104 Rf 1965	105 Db 1970	106 Sg 1974	107 Bh 1976	108 Hs 1984	109 Mt 1982										
LANTHANIDES		57 La 1839	58 Ce 1803	59 Pr 1885	60 Nd 1843	61 Pm 1947	62 Sm 1879	63 Eu 1896	64 Gd 1880	65 Tb 1843	66 Dy 1886	67 Ho 1878	68 Er 1843	69 Tm 1879	70 Yb 1878	71 Lu 1907		
ACTINIDES		89 Ac 1899	90 Th 1828	91 Pa 1917	92 U 1789	93 Np 1940	94 Pu 1940	95 Am 1945	96 Cm 1944	97 Bk 1950	98 Cf 1950	99 Es 1952	100 Fm 1953	101 Md 1955	102 No 1958	103 Lr 1961		

