The Periodic Table and Periodicity

Department of Chemistry
The Open University of Sri Lanka
Introduction

In this lesson, we will discuss the classification of elements in the Periodic Table. The Periodic Table helps us in the prediction of the properties of new compounds by comparison of the properties of known compounds. The modern Periodic Table is an arrangement of all the chemical elements in the order of increasing atomic number with elements having similar properties (i.e. of the same chemical family) in the same vertical column. In this session, we will learn about the group numbers, s, p, d and f blocks, and the electronic configuration. Some physical properties of elements such as atomic and ionic sizes/radii, ionization energy, electron affinity, electronegativity, melting and boiling points will be discussed in the next lesson. First of all we will briefly look at the history of the development of the Periodic Table.

1.1 Brief history of the development of the Periodic Table

In the early 19th century many new elements were being discovered and chemists were looking for similarities between these new elements and the existing elements.

Dobereiner (1829) suggested that elements could be grouped in three (triads), in which each member of the triad shows similar properties.

\[
e.g.: \quad \text{Lithium, sodium, potassium,} \\
\text{Calcium, strontium, barium} \\
\text{Chlorine, bromine, iodine}
\]

Newlands (1863) arranged the elements in the order of increasing relative atomic mass. He noticed that there was some similarity between every eighth elements (octet rule).
Moseley (1869) looked at the relationship between relative atomic mass and the density of an element. He then plotted a graph of atomic volume (mass of 1 mole of atoms divided by density) against the relative atomic mass for each element. The curve he obtained showed periodic variations.

Mendeleev/Mendeleef (1869) arranged the elements in order of increasing relative atomic mass but took into account the pattern of behaviour of the elements. He found that it was necessary to leave gaps in the table and said that these were for elements not known at that time. His table enabled him to predict the properties of the undiscovered elements. His work was proved correct by the accurate prediction of the properties of gallium and germanium. The Periodic Table we use today closely resembles that drawn up by Mendeleev.

A modification of the Periodic Table was made following the work of Rutherford and Moseley. It was realized that the elements should be arranged in the order of atomic number, i.e. the number of protons in the nucleus. In the Modern Periodic Table (Figure 1) the elements are arranged in the order of increasing atomic number with elements having similar properties in the same vertical column. Let us study about the Modern Periodic Table.

1.2 The Modern Periodic Table

Hydrogen, with the atomic number one, is the first element in the Periodic Table. Over one hundred elements are known and they are arranged in the order of increasing atomic number (Z). Elements are grouped (e.g. Group 1, 2, 3, ……, 17, 18) according to their chemical properties, which largely depend on the number of electrons in the outer/valence shell. The Periodic Table has vertical columns (18 Groups) and horizontal rows (7 Periods). The 1st, 2nd, 3rd, 4th, 5th and 6th periods have 2, 8, 8, 18, 18 and 32 elements, respectively. Majority of the elements are metals and there are about 17 non-metals placed at the right hand side of the Periodic Table. There are seven metalloids (B, Si, Ge, As, Sb, Te and At) which show
properties of both metals and non-metals. We will study about the groups and Group numbers now.

Figure 1: The Periodic Table with Group numbers and s, p, d and f blocks

Group number
The most recent Periodic Table assigns Group numbers (1 to 18) to the vertical columns in the Periodic Table. Thus, the elements Li, Na, K, Rb, Cs and Fr belong to Group 1 (formerly Group I or IA); Be, Mg, Ca, Sr, Ba and Ra belong to Group 2 (formerly Group II or IIA); Sc, Y, La and Ac belong to Group 3 (formerly Group IIIB); Fe, Ru and Os belong to Group 8; (formerly Group VIIIIB); B, Al, Ga, In and Tl belong to Group 13 (formerly Group III or IIIA); the halogens F, Cl, Br, I and At belongs to Group 17 (formerly Group VII or VIIA) and the noble gases He, Ne, Ar, Kr, Xe and Rn belong to Group 18 (formerly Group 0).

Blocks
The Periodic Table has four blocks (s, p, d and f) as s, p, d and f levels are being filled. The s-block consists of the elements in Groups 1 and 2. The elements in Groups 1 and 2 are also called alkali metals and alkaline earth metals respectively. The p-block consists of elements
in the Groups from 13 to 18. The \textbf{d-block} is located between s and p blocks and it consists of elements in the Groups from 3 to 12. The \textbf{f-block} has 28 elements and it is equally divided into \textbf{Lanthanides} \((Z = 58 \text{ to } 71)\) and \textbf{Actinides} \((Z = 90 \text{ to } 103)\). In most Periodic Tables, hydrogen is placed above lithium although it shows different properties to those of Li, Na, K and Cs. In this lesson, hydrogen will be considered separately.

\textbf{Hydrogen}

Hydrogen is the first element in the Periodic Table. Hydrogen atom has a very simple structure consisting of a nucleus with a single proton and one electron. Three isotopes of hydrogen are protium or proton \((H = ^1\text{H})\), deuterium or deuteron \((D = ^2\text{H})\) and tritium \((T = ^3\text{H})\). At room temperature, hydrogen is a diatomic, colourless and odourless gas.

In some Periodic Tables, hydrogen is placed above the Group 1 elements but sometimes it is placed above the Group 17 elements. \textit{It is best to have a separate place for hydrogen.}

\textbf{Q : Give reasons for placing hydrogen with the Group 17 elements.}

\textbf{A :} Hydrogen needs just one electron to attain the inert gas configuration \((1s^2)\). Halogens also need only one electron to achieve the inert gas configuration \((ns^2np^6)\). Hydrogen forms the hydride ion, \(H^-\) with electropositive metals. Halogens also form mono-negative anions, \(X^-\). Hydrogen \((H_2)\) is a diatomic gas similar to most halogens \((X_2)\).

\textbf{Activity}

1. Give reasons for placing hydrogen with the Group 1 elements.

The most important use of hydrogen is for the synthesis of ammonia \((\text{NH}_3)\) by the Haber process. It is also used to synthesize methyl alcohol from carbon monoxide \((\text{CO})\) in the Fischer-Tropsch synthesis. Hydrogen is also used for the conversion of vegetable fats and oils to give edible fats \((e.g.\) margarine). It is sometimes useful as a reducing agent in the

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extraction of metals, as a fuel (hydrogen powered cars), in oxy-hydrogen welding, and the production of organic compounds.

### s-Block Elements

The elements in the Groups 1 and 2 are called *s-block* elements.

#### Group 1 elements

The Group 1 elements (formerly Group I or IA) include lithium (Li), sodium (Na), potassium (K), rubidium (Rb), caesium (Cs) and francium (Fr). They are also called “alkali metals” and they conduct electricity and heat. Alkali metals are reactive and they are not found in nature as free metals. Sodium and potassium are relatively more abundant than the other metals in this group. The general valence electron configuration of these elements is ns\(^1\).

#### Group 2 elements

Group 2 (formerly Group II or IIA) elements include beryllium (Be), magnesium (Mg), calcium (Ca), strontium (Sr), barium (Ba) and radium (Ra). They are also called “alkaline earth metals”. The Group 2 elements are too reactive to occur as free elements in nature and all are powerful reducing agents. The general valence electron configuration of Group 2 elements is ns\(^2\).
**p-Block Elements**

The elements in the Groups 13 to 18 belong to *p-block*.

### Group 13 elements

This family (formerly Group III or IIIA) consists of boron (B), aluminium (Al), gallium (Ga), indium (In) and thallium (Tl). Boron is regarded as a *non-metallic* element with some metallic characteristics (*i.e.*, B is a semi-metal). The other elements are all metals. The general valence electron configuration of these elements can be written as $ns^2np^1$. All the elements show the oxidation state $+3$.

![Boron](boron26.png), [Aluminium](aluminium27.png), [Gallium](gallium28.png), [Indium](indium29.png), [Thallium](thallium30.png)

### Group 14 elements

The Group 14 elements (formerly Group IV or IVA) include carbon (C), silicon (Si), and germanium (Ge), tin (Sn) and lead (Pb). Carbon is a *non-metal*; silicon and germanium are *metalloids*; tin and lead are *metals*. The general valence electron configuration of Group 14 elements is $ns^2np^2$ and they exhibit a wide range of oxidation states from $-4$ to $+4$. The oxidation numbers of carbon in CH$_4$, diamond, CO and CO$_2$ are $-4$, $0$, $+2$ and $+4$. 

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respectively. Elemental carbon exists in three allotropic forms; allotropes are diamond, graphite and fullerene.

![Figure 2: Allotropes of carbon; diamond, graphite and fullerene (C_{60})^{31}](image)

Coal is still used as a fuel source. Carbon is also used as an electrode material. Carbon has three isotopes: $^{12}$C, $^{13}$C and $^{14}$C. Carbon-14 isotope is used in archaeological dating, for example, to determine the age of fossils. Graphite is a good lubricant and diamond is a precious material. Fullerenes have found various applications in the field of nanotechnology. Ultra-pure silicon is used in electronic industry. Polymeric silicones are used as oils and waxes.

![Elemental Carbon (C), Silicon (Si), Germanium (Ge), Tin (Sn), Lead (Pb)](image)

**Group 15 elements**

The Group 15 elements include nitrogen (N), phosphorus (P), arsenic (As), antimony (Sb) and bismuth (Bi). N and P are non-metals, As and Sb are metalloids whereas Bi has characteristic properties of a main group metal. The general valence electron configuration of Group 15 elements is $n{s^2}{p^3}$ and they exhibit a range of oxidation states from $-3$ to $+5$. The oxidation numbers of nitrogen in NH$_3$, N$_2$, NF$_3$ and NO$_3^-$ are $-3$, $0$, $+3$ and $+5$, respectively. Air contains 78% of N$_2$ (by volume). Nitrogen is obtained by the fractional distillation of liquefied air. Conversion of inert nitrogen gas into useful nitrogen compounds is called “Nitrogen fixation”. NH$_3$ is commercially produced in the Haber synthesis by reacting N$_2$ and H$_2$ in the presence of an iron catalyst at 500°C and at a pressure of 300 atm. Nitric acid is manufactured by oxidising ammonia using air.
**Group 16 elements**
This family consists of the elements: **oxygen** (O), **sulfur** (S), **selenium** (Se), **tellurium** (Te) and **polonium** (Po). The Group 16 elements are also known as “**chalcogens**”. Polonium is the only true metal; Te is a semi-metal; oxygen, sulfur and selenium are non-metals. The general valence electron configuration of Group 16 elements is \( ns^2np^4 \) and they exhibit a wide range of oxidation states from −2 to +6. The oxidation numbers of sulphur in \( \text{H}_2\text{S}, \text{S}_8, \text{SO}_2 \) and \( \text{SO}_3 \) are −2, 0, +4 and +6, respectively. In general, the reactivity of these elements decreases in the following order \( \text{O} > \text{S} > \text{Se} > \text{Te} \). Oxygen (\( \text{O}_2 \)) is the second most abundant gas in the earth’s atmosphere and it is obtained from liquid air by fractional distillation. Photolysis of the oxygen generates two oxygen radicals which combine with another oxygen molecule to form ozone (\( \text{O}_3 \)).

**Group 17 elements**
This family consists of **fluorine** (F), **chlorine** (Cl), **bromine** (Br), **iodine** (I) and **astatine** (At). Halogens are non-metals and they exist as **diatomic** molecules. At room temperature, fluorine and chlorine are yellow green gasses; bromine is a red liquid and iodine a black solid. Solutions of iodine are purple in colour. Fluorine is the most electronegative element and a powerful oxidizing agent. Fluorine is very reactive and forms compounds with most elements.

The general valence electron configuration of an element is \( ns^2np^5 \). They achieve the next noble gas electron configuration by gaining one electron from a metal or sharing an electron with a non-metal. Thus, all elements show the oxidation number −1. Positive oxidation...
numbers such as +1, +3, +5 and +7 occur for chlorine, bromine and iodine, mainly in oxyanions and inter-halogen compounds. The bond dissociation energies, D(X-X), of the halogen molecules decrease in the following order. D(Cl-Cl) > D(Br-Br) > D(F-F) ≈ D(I-I)

**Group 18 elements**
The Group 18 elements are normally called the **noble gases** and they include **helium** (He), **neon** (Ne), **argon** (Ar), **krypton** (Kr), **xenon** (Xe) and **radon** (Rn). They are called noble gases because they are chemically not very reactive. They are also called **rare gases** as they are found only in very small quantities in the atmosphere and in the earth’s crust. All are **monatomic**, colourless and odourless gases. Radon is radioactive. The general valence electron configuration of Ne, Ar, Kr, Xe and Rn is ns²np⁶ (the total number of electrons in the outermost shell is 8). The electron configuration of He is 1s². In all noble gases, the valence shell has the maximum number of electrons, thus atoms do not combine with one another (to form diatomic gas molecules). They also do not combine readily with other elements. This is the reason why they are referred to as 'inert' gases. But since 1960s fluorides and oxides of Xe are known.
**d-Block Elements (Groups 3 to 12)**

All d-elements are metals and they are generally called transition metals. These elements are obtained by the filling up of d-levels. Each metal consists of a positively charged lattice and a pool of free electrons. All d-elements are good heat and electrical conductors.

<table>
<thead>
<tr>
<th>3d</th>
<th>4d</th>
<th>5d</th>
<th>6d</th>
</tr>
</thead>
<tbody>
<tr>
<td>3</td>
<td>39</td>
<td>57</td>
<td>89</td>
</tr>
<tr>
<td>Sc</td>
<td>Y</td>
<td>La</td>
<td>Ac</td>
</tr>
<tr>
<td>Ti</td>
<td>Zr</td>
<td>Hf</td>
<td>Rf</td>
</tr>
<tr>
<td>V</td>
<td>Nb</td>
<td>Ta</td>
<td>104</td>
</tr>
<tr>
<td>Cr</td>
<td>Mo</td>
<td>W</td>
<td>105</td>
</tr>
<tr>
<td>Mn</td>
<td>Tc</td>
<td>Re</td>
<td>Db</td>
</tr>
<tr>
<td>Fe</td>
<td>Ru</td>
<td>Os</td>
<td>Sg</td>
</tr>
<tr>
<td>Co</td>
<td>Rh</td>
<td>Ir</td>
<td>Bh</td>
</tr>
<tr>
<td>Ni</td>
<td>Pd</td>
<td>Pt</td>
<td>Hs</td>
</tr>
<tr>
<td>Cu</td>
<td>Ag</td>
<td>Au</td>
<td>Mt</td>
</tr>
<tr>
<td>Zn</td>
<td>Cd</td>
<td>Hg</td>
<td>Ds</td>
</tr>
</tbody>
</table>

Most of the d-elements show variable oxidation states. d-Block elements have higher melting and boiling points than s- and p-block elements. In the 3d-series, the melting point gradually increases from Sc to V and then decreases from Fe to Zn. d-Block elements show higher densities compared to s- and p-block elements. In the 3d-series, the density increases from Sc to Cu. The d-block elements are less electropositive than the Group 1 metals. In the 3d-series, the electronegativity tends to increase slightly from Sc to Cu. Majority of the transition metal complexes are coloured. d-Block elements show variable oxidation states and their metal complexes play an important role as catalysts. Zn, Cd and Hg are not considered as transition metals. Since scandium does not form stable Sc⁺ (d³) or Sc²⁺ (d¹) ions, in many text books, it is not considered a transition metal.

**Activity**

2. Why is Zn not considered as a transition metal?

**f-Block Elements**

f-Block contains 28 elements and it is placed at the bottom of the Periodic Table. These elements are obtained by the filling up of f-levels, see section on electron configuration for more details.
f-Block has two horizontal rows: the elements in the first row are called Lanthanides (or Lanthanoids) whilst the elements in the second row are called Actinides (or actinoids). However, the lanthanide elements (i.e. elements after lanthanum (La) having atomic numbers 57 to 71) should be placed at the period 6 between the elements La and Hf. Similarly, the actinide elements (i.e. elements after actinium having atomic numbers 90 to 103) should be placed at the period 7. Electron configurations of elements can be used to group elements into respective blocks. Reactivity of elements, metallic and non-metallic properties of elements, and ionization energies of elements can be easily predicted by considering the electron configuration of an element.

**Electron configuration**
The electrons are found in different energy levels or shells (K, L, M, N, etc). Each shell consists of a number of sub-shells (s, p, d, f, etc). The number of orbitals found in s, p, d and f sub-shells are 1, 3, 5 and 7, respectively. The maximum number of electrons found in an orbital is two. Distribution of orbitals among the first four shells is given in Table 1.

<table>
<thead>
<tr>
<th>n</th>
<th>Shell</th>
<th>Orbitals</th>
<th>Total number of orbitals</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>K</td>
<td>1s</td>
<td>1</td>
</tr>
<tr>
<td>2</td>
<td>L</td>
<td>2s, 2p&lt;sub&gt;x&lt;/sub&gt;, 2p&lt;sub&gt;y&lt;/sub&gt;, 2p&lt;sub&gt;z&lt;/sub&gt;</td>
<td>4</td>
</tr>
<tr>
<td>3</td>
<td>M</td>
<td>3s, 3p&lt;sub&gt;x&lt;/sub&gt;, 3p&lt;sub&gt;y&lt;/sub&gt;, 3p&lt;sub&gt;z&lt;/sub&gt;, 3d&lt;sub&gt;xz&lt;/sub&gt;, 3d&lt;sub&gt;xy&lt;/sub&gt;, 3d&lt;sub&gt;x&lt;sup&gt;2&lt;/sup&gt;-y&lt;sup&gt;2&lt;/sup&gt;&lt;/sub&gt;, 3d&lt;sub&gt;z&lt;sup&gt;2&lt;/sup&gt;&lt;/sub&gt;</td>
<td>9</td>
</tr>
<tr>
<td>4</td>
<td>N</td>
<td>4s, 4p&lt;sub&gt;x&lt;/sub&gt;, 4p&lt;sub&gt;y&lt;/sub&gt;, 4p&lt;sub&gt;z&lt;/sub&gt;, 4d&lt;sub&gt;xz&lt;/sub&gt;, 4d&lt;sub&gt;xy&lt;/sub&gt;, 4d&lt;sub&gt;x&lt;sup&gt;2&lt;/sup&gt;-y&lt;sup&gt;2&lt;/sup&gt;&lt;/sub&gt;, 4d&lt;sub&gt;z&lt;sup&gt;2&lt;/sup&gt;&lt;/sub&gt;, Seven 4f orbitals</td>
<td>16</td>
</tr>
</tbody>
</table>
Q: What is the order of filling of energy levels?

A: 1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d, 5p, 6s, 4f, 5d, 6p, 7s, etc.

In terms of energy: 1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < 5s ..........

The elements of the Periodic Table can be further classified into four main categories depending on the electronic configurations of the atoms. They are:

1. Noble gases (Group 18)
2. Representative elements (s- and p-blocks)
3. Transition elements (d-block)
4. Inner-transition elements (f-block)

**Noble gases**

These elements are classified under Group 18 of the Periodic Table. Except helium (which has two electrons in the 1s orbital; thus the valence electron configuration is 1s²), all the other elements have the valence electron configuration ns²np⁶. These elements have a filled valence shell thus they are inert. All are monatomic gases.

Q: What are the electron configurations of Ar and K?

A: The electron configuration of Ar is 1s²2s²2p⁶3s²3p⁶, which can be represented as [Ar].

The electron configuration of K is 1s²2s²2p⁶3s²3p⁶4s¹.

Thus, the electron configuration of K can be written as [Ar]4s¹.

**Note** that the electron configuration of K⁺ is 1s²2s²2p⁶3s²3p⁶ = [Ar]; i.e. K⁺ and Ar have the same number of electrons (isoelectronic). Similarly, Ca²⁺, Ga³⁺, P³⁻, S²⁻ and Cl⁻ are isoelectronic with Ar.
**Representative elements**

These are the elements with the valence electron configurations from \(ns^1\), \(ns^2\), \(ns^2np^1\), \(ns^2np^2\), \(ns^2np^3\), \(ns^2np^4\) to \(ns^2np^5\). These elements belong to s- and p-blocks.

Q : What is the electron configuration of P?

A : The electron configuration of P is \(1s^22s^22p^63s^23p^3\) = \([\text{Ne}]3s^23p^3\).

Note that the electron configuration of \(P^{3-}\) is \(1s^22s^22p^63s^23p^6\) = \([\text{Ar}]\).

**Transition elements**

*The atoms or the stable ions* of these elements have *incompletely filled d-shells* thus they are called “transition metals”. In general they can have either one or two electrons in the last s-shell and one to ten electrons in the penultimate (i.e. one before the last) d-shell. The valence electron configuration can be written as \(ns^1\) or \(2(n-1)d\) to \(10\). The valence electron configuration of a d-block element can also be represented as \((n+1)s^2nd^m\) or \((n+1)s^1nd^m\) where \(n = 3, 4, 5\) or 6 and \(m = 1, 2, 3, \ldots\) or 10.

The first transition series (3d-series) starts with the Group 3 element scandium Sc (with the atomic number, Z of 21) which has the electron configuration \(1s^22s^22p^63s^23p^64s^23d^1\) and ends up with the Group 11 element Cu (Z = 29) with the electron configuration \(1s^22s^22p^63s^23p^64s^13d^{10}\). Although copper in the element form has a filled d shell, it forms the Cu\(^{2+}\) ion with the electron configuration \(1s^22s^22p^63s^23p^63d^9\), and it is a transition metal.

Note that for free atom, the energies of the 3d and 4s orbitals are very close (3d > 4s). In metal cations 3d orbitals are much more stabilised than the 4s orbitals (i.e.3d < 4s), thus 4s electrons are lost first. Therefore the electron configuration of Ti\(^{2+}\) and V\(^{3+}\) ions can be written as \([\text{Ar}]3d^2\).

Q : What is the electron configuration of the Cu\(^+\) ion?

A : The electron configuration of the Cu\(^+\) ion is \(1s^22s^22p^63s^23p^63d^{10}\) = \([\text{Ar}]3d^{10}\)
**Inner-transition elements**

These elements are obtained by the filling up of f-levels. For example, the fourteen elements from Ce (Z = 58) to Yb (Z = 70) are obtained by filling of the 4f-level. In this series, called *Lanthanides*, the filling of electrons is rather *irregular* in the sense that the 5d energy level is occasionally filled in preference to the 4f-level. Their electron configurations can be written as [Xe]\(^{4f^n5d^16s^2}\) \((m = 1-14)\). The elements obtained by similar filling up of the 5f level are called *Actinides*.

Q : What is the electron configuration of Ce?

A : The electron configuration of the Ce is [Xe]\(^{4f^15d^16s^2}\).

Note that in the free atom, the order of filling of energy levels is 6s < 5d < 4f.

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**Activity**

3. (a) Write the electron configurations of the following atoms/ions.

   (i) K   (ii) Kr   (iii) Cr   (iv) Fe\(^{2+}\)   (v) Cl\(^-\)   (vi) Si

   (b) Write the valence electron configurations of the following atoms/ions.

   (i) Cs\(^+\)   (ii) Ir\(^{3+}\)   (iii) Sr   (iv) Mn\(^{7+}\)   (v) S\(^{2-}\)
2. Periodicity

Introduction
We have studied the classification of the elements in the Periodic Table. Let us now consider some physical properties of the elements: atomic size, atomic radius, ionic radius, ionization energy, electron affinity, electronegativity, melting and boiling points etc.

2.1 Atomic and ionic sizes
A few elements exist as atoms at ambient temperatures. Most elements form neutral molecules (e.g. N₂, H₂, etc.) or ions. Addition or removal of an electron(s) to or from an atom generates ions. The size of an atom depends on the effective volume of the outer electrons. Atomic sizes of elements are important because properties such as boiling point, melting point, ionization energy and electronegativity depend on their sizes. The size of an atom/ion can be estimated from the measured distance between the nuclei of two neighbouring atom/ion by making the assumption that this distance is equal to the sum of the radii of the two atom/ion. Because this distance depends on the type of bond between the atoms, there are few types of radii, namely, atomic, covalent, ionic, metallic radii etc. Each radius is calculated for a particular type of bond.

2.2 Atomic radii
The radius of an atom can be considered as the distance of closest approach to another identical atom. This definition does not mention whether the atoms are chemically bonded or not, thus radius depends on the type of bond between the atoms/ions. Therefore, we use covalent radius for covalent compounds, i.e. the equilibrium distance between two covalently bonded atoms. In free metals we use metallic radius, i.e. the half distance between the neighbouring atoms in the metal.

Several experimental methods are available for measuring inter nuclear distances and hence for calculating the radii of atoms. For example, in the Cl₂ molecule, which is formed by sharing a pair of electrons between two chlorine atoms, the Cl–Cl distance is found to be
0.198 nm. Half of this distance (0.099 nm) will be the covalent radius of chlorine. In diamond, the C–C bond distance is found to be 0.154 nm and the same value is found for the C–C bond distance in many other aliphatic compounds. Hence the covalent radius of carbon is 0.077 nm. In free metals, each atom has many neighbours and the bonds between the atoms are known as “metallic bond”. Metallic radii are usually larger than covalent bond radii.

The atomic radius increases as the number of energy levels occupied by electrons increases. Hence the elements within a group in the Periodic Table would show a gradual increase in atomic radius as the atomic number increases. The atomic radii of the alkali metals (Li to Cs) are given below for comparison.

<table>
<thead>
<tr>
<th>Element</th>
<th>Li</th>
<th>Na</th>
<th>K</th>
<th>Rb</th>
<th>Cs</th>
</tr>
</thead>
<tbody>
<tr>
<td>Atomic radius (nm)</td>
<td>0.123</td>
<td>0.157</td>
<td>0.203</td>
<td>0.216</td>
<td>0.235</td>
</tr>
</tbody>
</table>

**Activity**

4. Compare the atomic radii of the alkaline earth metals.

Within a period there is a general decrease in the radii of the atoms from left to right of the Periodic Table. This decrease is mainly due to the increase of positive charge in the nucleus (i.e. the number of protons). The decrease in atomic radius as the atomic number increases from Li to F is given below for comparison.

<table>
<thead>
<tr>
<th>Element</th>
<th>Li</th>
<th>Be</th>
<th>B</th>
<th>C</th>
<th>N</th>
<th>O</th>
<th>F</th>
</tr>
</thead>
<tbody>
<tr>
<td>Atomic radius/nm</td>
<td>0.123</td>
<td>0.106</td>
<td>0.088</td>
<td>0.077</td>
<td>0.074</td>
<td>0.073</td>
<td>0.064</td>
</tr>
</tbody>
</table>

Across a period, from one element to the next, each additional electron added occupies the same principal energy level. Thus, the effective positive charge increases from left to right across a period but the increase in positive charge is not neutralized by increasing shielding by shielding electrons (see section 2.5). Hence, the outermost electrons of an atom are drawn closer to the nucleus resulting in a smaller atomic radius. Note that atomic radii of noble gases may be greater than those of the other elements in the same period.
The variation in covalent radii of the elements from hydrogen to potassium is given in Fig. 3.

![Covalent radii graph](image)

**Figure 3**: The variation in covalent radii of the elements from H to K

### Activity

5. Compare the atomic radii of the elements in the 3rd period (Na to Cl).

### 2.3 Ionic radii

When atoms lose or gain electrons, ions are formed. The radii of such ions (*i.e.* ionic radii) can be determined from the distances between the ions in a crystal. For example, the ionic radii of the sodium ion (Na⁺) and chloride ion (Cl⁻) can be obtained from the inter-nuclear distances in sodium chloride crystals. X-ray diffractometer is used to determine the molecular crystal structure of a compound and atomic/ionic distances can be easily determined.

Positive ions (cations) are formed by the loss of electron(s) and are always smaller than the parent atom. For example, the atomic radius of vanadium is 0.131 nm whereas the ionic radii of V²⁺, V³⁺ and V⁴⁺ are 0.088, 0.074 and 0.060 nm, respectively.

Q : Explain the above observation.

A : The size of an atom or ion decreases as the net positive charge on the particle increases.

The net positive charge present in a positive ion attracts the electrons more strongly towards the nucleus; this would lead to a contraction of the electron cloud. Thus, ionic radii decrease as the positive charge increases.
Addition of an electron to an atom generates a negatively charged ion (or anion). The incoming electron enters the highest energy level without an increase in nuclear charge (i.e. there is no change in the number of proton in the nucleus). There is electron-electron repulsion with an expansion of the electron cloud. Therefore, the size of the negatively charged ion increases over that of the atom. For example, the atomic radius of chlorine atom is 0.099 nm whereas the radius of the Cl\(^-\) ion is 0.181 nm.

Addition of an electron to an atom generates a negatively charged ion (or anion). The incoming electron enters the highest energy level without an increase in nuclear charge (i.e. there is no change in the number of proton in the nucleus). There is electron-electron repulsion with an expansion of the electron cloud. Therefore, the size of the negatively charged ion increases over that of the atom. For example, the atomic radius of chlorine atom is 0.099 nm whereas the radius of the Cl\(^-\) ion is 0.181 nm.

**Activity**

6. The atomic radius of Li is 0.123 nm whereas the radius of the Li\(^+\) ion is 0.068 nm. Explain this observation.

**Activity**

7. The atomic radius of iodine is 0.128 nm whereas the radius of the I\(^-\) ion is 0.219 nm. Explain.

### 2.4 Metallic and non-metallic character

The tendency of an element to lose electrons to generate a positively charged ion (cation) is called its metallic character (which is related to electro-positivity). The tendency of an element to gain electrons to form negatively charged ions (anions) is called its non-metallic character. The metallic character of the elements increases as we go down a group; for example metallic character increases from Li to Cs as the size of the atoms increases when we go down a group. As a result, the electrons of the outermost shell can be easily removed. We can also say that the non-metallic character decreases from Li to Cs.

**Activity**

8. (a) Arrange the following elements in the order of increasing metallic character.

   As, N, P and Sb

   (b) Arrange the following elements in the order of increasing non-metallic character.

   Be, Ca, Mg and Sr
Across a period, from left to right, the metallic character of the elements decreases and the non-metallic character increases. The reasons are (i) the atomic radius decreases from left to right and (ii) the effective nuclear charge increases from left to right. For example, among the elements in the second period, Li is the most metallic element and F is the most non-metallic element. The noble gases do not exhibit any metallic or non-metallic character as they have no tendency to lose or gain electrons.

2.5 Ionization Energy

The **first ionization energy** of an atom is the energy required to remove an electron from an atom in the gaseous state to an infinitely large distance. The process can be written as

\[
\text{M}(g) \rightarrow \text{M}^+(g) + e^- 
\]

Since ionization is a process common to many chemical reactions, it is an important factor in interpreting reactivity of many elements.

The factors that affect the ionization energy of an atom are given below.

(a) **Atomic radius**: As the atomic radius increases, the ionization energy becomes smaller.

(b) **The effective nuclear charge**: When the effective charge of the nucleus increases, the ionization energy becomes larger.

(c) **Shielding by inner electrons**: When an atom has more inner shells between the nucleus and the outer most electrons, the outer most electrons can be easily removed. The reasons are (i) inner-shell electrons (shielding electrons) repel outer most electrons and (ii) inner-shell electrons shield the outer most electrons from the nucleus; thus the effect of effective nuclear charge decreases as shielding increases (shielding effect). Therefore, the ionization energy becomes smaller with the increase of shielding. Shielding effect decreases in the following order \( s > p > d > f \).

The **first ionization energy** refers to the removal of the most loosely bound electron from an isolated atom in the gaseous state. The second, third, fourth ... ionization energies refer to the removal of the second, third, fourth ......electrons and so on. Since the energy required to remove an electron depends on how strongly it is bound to the nucleus of an atom, the ionization energy will be lowest when the effective nuclear charge is smallest and when the distance separating the electron from the nucleus is largest. The second ionization energy
refers to the energy required to remove one electron from the gaseous mono-cation, $M^+(g)$, as shown below.

$$M^+(g) \longrightarrow M^{2+}(g) + e^-$$

The third ionization energy refers to the energy required to remove one electron from the gaseous dication, $M^{2+}(g)$, as shown below.

$$M^{2+}(g) \longrightarrow M^{3+}(g) + e^-$$

*Ionization energy generally increases from left to right* as (i) the atomic radius decreases from left to right and (ii) the effective nuclear charge increases from left to right of the periodic table. Shielding effect is very small as all outermost electrons are in the same shell.

However, there are several irregularities; *i.e.* some elements have ionization energy higher than that of the next element in a period. This slight deviation can be explained by considering the *electron configurations of the atoms and the ions formed*. It is known that completely filled and half-filled shells and sub shells show extra stability. Thus elements with $s^2$, $s^2p^6$ (completely filled) and $s^2p^3$ (half filled) configurations show higher ionization energies. It is also found that when the resulting cation has an electron configuration with a completely filled or half-filled shell then the ionization energy is found to be low. Figure 4 shows the first ionization energies of the first fifty four elements in the Periodic Table.

*Figure 4: The first ionization energies of the first fifty four elements*
It is clear from Figure 4 that the first ionization energy increases across a period with some irregularities. For example, the first ionization energy of beryllium is greater than that of boron.

Q: How would you explain the above observation?
A: The valence electron configuration of boron is 2s²2p¹. The valence electron configuration of beryllium is 2s², which has a filled s sub shell. Thus removal of an electron from beryllium is much more difficult than boron. Removal of one electron from boron gives a filled s sub shell (2s²) with extra stability; therefore, ionization energy of boron is smaller than that of beryllium.

The first ionization energy of noble gasses is much higher when compared with those of other elements in the same period. Ionization energy is expressed in kJ mol⁻¹ or eV where 1 eV = 96.85 kJ mol⁻¹.

Ionization energy decreases as you go down a group. In this case, all elements have the same number of valence electrons. As we go down a group the valence electrons are held less tightly as they are shielded (and also repelled) by the electrons in the inner shells. Thus, removal of an electron from outermost orbital is much easier. As a result, the first ionization energy decreases as you go down a group. The first ionization energies of the alkali metals (Li to Cs) are given below for comparison.

<table>
<thead>
<tr>
<th>Element</th>
<th>Li</th>
<th>Na</th>
<th>K</th>
<th>Rb</th>
<th>Cs</th>
</tr>
</thead>
<tbody>
<tr>
<td>1st Ionization Energy/kJmol⁻¹nm</td>
<td>520</td>
<td>513</td>
<td>419</td>
<td>400</td>
<td>380</td>
</tr>
</tbody>
</table>

Activity
9. The first ionization energy of magnesium is greater than that of aluminium. Explain

2.6 Electron affinity

Electron affinity is defined as the amount of energy released when an electron is added to a neutral gaseous atom. It is the energy released during this reduction process,

\[ \text{X(g)} + \text{e}^- \rightarrow \text{X}^-(\text{g}) \]
This can be considered as the reverse of the ionization of $X^-(g)$ ion.

$$X^-(g) \longrightarrow X(g) + e^-$$

Electron affinity increases (increase in the negative value or the energy liberated) from left to right of the Periodic Table as (i) the effective nuclear charge increases and (ii) the size of the atom decreases, from left to right. For most elements, energy is liberated when electron is added to a neutral gaseous atom, and the electron affinity therefore takes a negative value. However, there are some elements with positive electron affinity values i.e. energy should be supplied from outside to force an electron to combine with the gaseous atom, to form an ion. Affinity for the second electron needs a large energy because the second electron must be forced on against the net negative charge of the ion. This is illustrated by the following example:

$$\text{O}(g) + e^- \longrightarrow \text{O}^-(g) \quad E = -140 \text{ kJ mol}^{-1} \quad \text{(energy liberated)}$$

$$\text{O}^-(g) + e^- \longrightarrow \text{O}^{2-}(g) \quad E = +700 \text{ kJ mol}^{-1} \quad \text{(energy supplied)}$$

**Activity**

10. Electron affinity of chlorine is negative but that of argon is positive, explain.

## 2.7 Electronegativity

The electronegativity of an atom is a measure of the tendency of the atom to attract electrons when it is covalently bonded to other atom(s). Generally small atoms attract electrons more strongly than larger ones and are therefore more negative. Like ionization energy and electron affinity, electronegativity describes the tendency of an atom to attract electrons but here we consider a particular bonding situation. Unlike ionization energy and electron affinity, electronegativity cannot be measured directly.

There are several ways of expressing the electronegativity of elements. For example, in the Pauling’s method the most electronegative element fluorine (F) is given an arbitrary value of 4.0. The electronegativity of other atoms is defined with reference to the value of fluorine. Electronegativity values for some selected elements are given below.

<table>
<thead>
<tr>
<th>Element</th>
<th>F</th>
<th>O</th>
<th>Cl</th>
<th>N</th>
<th>C</th>
<th>H</th>
<th>Li</th>
</tr>
</thead>
<tbody>
<tr>
<td>Electronegativity</td>
<td>4.0</td>
<td>3.5</td>
<td>3.0</td>
<td>3.0</td>
<td>2.5</td>
<td>2.1</td>
<td>1.0</td>
</tr>
</tbody>
</table>
Electronegativity is a useful quantity for predicting bond type, dipole moments and bond energies. For example, covalent bonds (i.e. formed by sharing of electrons between atoms) are formed by elements which have similar electronegativity values and ionic bonds (i.e. transfer of electrons from one to the other) are formed between elements with a large electronegativity difference.

The electronegativity of elements increase sharply across a period of the s and p block elements (for example from Li to F). This can be explained on the basis of the increased effective nuclear charge across a period. (Refer section 2.6 to see how the electron affinity varies across a period). The electronegativity generally decreases when going down a group. Figure 5 shows the variation of electronegativity across the third period.

![Figure 5: The variation of electronegativity across the third period](image)

**Activity**

11. Arrange the following elements in the order of increasing electronegativity.

(a) H, Li, C, N, O, F  (b) Be, Mg, Ca, Ba  (c) F, Cl, Br, I

**2.8 Melting and boiling points**

The melting point is defined as the temperature at which pure solid is in equilibrium with pure liquid. This temperature depends on the strength of the forces holding the
particles together in the solid and the extent to which these must be broken to form a liquid. The melting points you normally find in textbooks are usually measured at atmospheric pressure.

Within a group, there is a gradual change of melting points. As given in Table 2, the melting points of the alkali metals decrease when you go down the group and in the case of the halogens the melting points increase when you go down the group. In the case of alkali metals, the atoms are held together by weak metallic bonding. The smaller Li atoms pack more compactly in the solid and the metallic bonds in Li are stronger than that in Na, which is a larger atom. This is the reason for the gradual decrease in the melting points of alkali metals as you go down the group. In contrast halogens are held together by weak van der Waals forces. Thus halogens have relatively low melting points.

However, since structural changes are abrupt in going along a period, the melting point could change drastically when going from one element to the other across a period. As an example let us consider the adjacent element carbon and nitrogen; carbon is a solid at room temperature with a very high melting point whereas nitrogen is a gas with a very low melting point.

Table 2: Melting and boiling points of elements in Groups 1 and 17

<table>
<thead>
<tr>
<th>Element</th>
<th>m.pt./°C</th>
<th>b.pt./°C</th>
</tr>
</thead>
<tbody>
<tr>
<td>Li</td>
<td>181</td>
<td>1331</td>
</tr>
<tr>
<td>Na</td>
<td>98</td>
<td>890</td>
</tr>
<tr>
<td>K</td>
<td>64</td>
<td>766</td>
</tr>
<tr>
<td>Rb</td>
<td>39</td>
<td>701</td>
</tr>
<tr>
<td>Cs</td>
<td>29</td>
<td>685</td>
</tr>
<tr>
<td>F</td>
<td>-220</td>
<td>-188</td>
</tr>
<tr>
<td>Cl</td>
<td>-101</td>
<td>-34</td>
</tr>
<tr>
<td>Br</td>
<td>-7</td>
<td>58</td>
</tr>
<tr>
<td>I</td>
<td>114</td>
<td>183</td>
</tr>
</tbody>
</table>

Activity

12. Discuss the variation of melting points in Group 2 and Group 16 elements.
The variation in melting points of the elements from hydrogen to rubidium is given in Fig. 6.

![Figure 6: The variation in melting points of the elements from H to Rb](image)

**Boiling Points**

The periodic trends of boiling points are similar to those of the melting point (see Table 2). In the case of metals the process of boiling requires complete breaking of bonds since there is little or no bonding of atoms in the vapour phase. Appreciable metallic bonding exists in the liquid phase and these bonds must be broken during vaporization. Thus a large amount of thermal energy must be supplied to vaporize hence the boiling points are high. There is also a large deference between melting and boiling points. Non-metals such as N$_2$ and Cl$_2$ require a small energy to break up the weak van der Waals forces thus the boiling points are also relatively low. In contrast, non-metals such as C and Si which have giant covalent structures show high boiling points.

**2.9 Oxidation states**

There is a simple periodic relationship between the common oxidation states (numbers) of elements and their electron configuration. The oxidation number of any free element is zero. Fluorine shows only one oxidation state, −1.

Let us consider the s- and p-block elements. The maximum positive oxidation number shown by any representative element is equal to the total number of s- and p-electrons in the valance shell. For example, the Group 13 elements with the $s^2p^1$ configuration would have the maximum oxidation number of +3. The elements in the Groups 1 and 2 show only one positive oxidation state, e.g. Na$^+$ (+1) and Ca$^{2+}$ (+2).
Q : What are the maximum oxidation numbers of Chromium and Chlorine?
A : +6 and +7 respectively.

The common oxidation states for each Group of elements from Li to Ne generally correspond to the number of electrons gained or lost in order to achieve the noble gas configuration. An element with the \( n^{2}np^{4} \) valence electron configuration would gain two electrons from another element and achieve the octet configuration; thus, oxygen would show an oxidation state of \(-2\). Thus, the oxide of aluminium is \( \text{Al}_2\text{O}_3 \) and the formula of barium chloride is \( \text{BaCl}_2 \).

Transition elements generally have varying numbers of \( d \)-electrons and they show several oxidation states.

Q : What are the formulae of oxides of Mn?
A : \( \text{MnO}, \text{Mn}_2\text{O}_3, \text{MnO}_2, \text{MnO}_3, \text{Mn}_2\text{O}_7, \text{and MnO}_4^- \).

The oxidation numbers of Mn in \( \text{MnO}, \text{Mn}_2\text{O}_3, \text{MnO}_2, \text{MnO}_3, \text{Mn}_2\text{O}_7, \text{and MnO}_4^- \) are +2, +3, +4, +6, +7 and +7, respectively.

Activity

13. Determine the common oxidation number(s) of Groups 1, 2 and 16 using their electron configuration of representative element.

2.10 Diagonal relationships

There are certain irregularities related to the positions of many lighter elements. For example, the Group 2 element beryllium is slightly different from the rest of the members of the group; it resembles the Group 13 element aluminium in many ways. Another example is the similarity in the properties of lithium and magnesium. The oxidation states of these elements correspond to their group numbers, but the nature of
their compounds, their acidic and basic characteristics are related in a diagonal fashion. This is called diagonal relationship (relationship along the diagonal). This relationship can be explained in terms of ionic potential (equals to ratio of ionic charge/ionic radius). This term is really a measure of the charge density on the ion.

\[
\text{Ionic potential for } \text{Be}^{2+} = \frac{\text{ionic charge (q)}}{\text{ionic radius (r)}} = \frac{2}{0.038} = 52.6
\]

\[
\text{ Ionic potential for } \text{Al}^{3+} = \frac{3}{0.052} = 57.6
\]

Since the ionic potentials for the two ions given above are similar, the properties such as hydration energy, acid-base character and polarizing power of the elements are expected to be similar.

**Summary**

- The Periodic Table is an arrangement of all the chemical elements in the order of increasing atomic number with elements having similar properties in the same vertical column.
- The Periodic Table has eighteen groups and seven periods. The 1\(^{st}\), 2\(^{nd}\), 3\(^{rd}\), 4\(^{th}\), 5\(^{th}\) and 6\(^{th}\) periods have 2, 8, 8, 18, 18 and 32 elements, respectively.
- The Periodic Table has four blocks (s, p, d and f) due to the filling of s, p, d and f levels.
- Atomic sizes of elements are important because properties such as boiling point, melting point, ionization energy and electronegativity depend on their sizes.
- The atomic radius increases as the number of energy levels occupied by electrons increases.
- Within a period, from left to right, there is a general decrease in the radii of the elements.
- The factors that affect the ionization energy of an atom are atomic radius, effective nuclear charge, and shielding and repulsion by inner electrons.
- Generally, ionization energy increases from left to right of the Periodic Table and ionization energy decreases as you go down a group.
- Electron affinity is defined as the amount of energy released when an electron is added to a neutral gaseous atom. Electron affinity increases (increase in the negative value or the energy liberated) from left to right of the Periodic Table.
• The electronegativity of an atom is a measure of the tendency of the atom to attract electrons when it is covalently bonded to other atom(s). The electronegativity generally decreases when going down a group.

**Learning Outcomes**

At the end of the session you should be able to

• explain briefly the history of the development of the Periodic Table
• discuss the classification of elements into s, p, d and f-blocks
• write the electron configuration of the elements in the s, p, d and f-blocks
• discuss the periodic trends in atomic size, atomic radius, and ionic radius of elements.
• discuss the metallic and non-metallic character of elements
• define ionization energy of elements
• explain trends within groups and periods
• discuss the electron affinity, electronegativity, melting and boiling points
• comment on elements which show diagonal relationships

**Activity**

14. Define the term “electronegativity”.

15. Account for the difference in the ionization energies between the following successive pairs of elements in the Periodic Table. The values of ionization energy (in kJ mol⁻¹) are given in parentheses.
   a) He (2372) and Li (520)
   b) Li (520) and Be (900)
   c) Be (900) and B (800)
   d) N (1400) and O (1300)

16. There is a marked difference between the values for the fifth and the sixth ionization energies of nitrogen. Explain.

17. Define the first electron affinity of an atom.

18. Discuss the variation of electron affinity in the Periodic Table.

19. What are the formulae of common oxides of Cr?

20. What are the formulae of common chlorides of vanadium?
Answer Guide to Activities

1. The reasons for placing hydrogen with the Group 1 elements are (i) hydrogen has one electron in the outermost shell, (ii) hydrogen forms a mono-positive ion, $H^+$, (iii) like Group 1 metals, it gets discharged at the cathode.

2. Zn is not considered as a transition metal because (i) its d-shell is completely filled ($d^{10}$), (ii) it does not form coloured compounds, and (iii) it does not show variable oxidation states ($Zn^{2+}$).

3. (a) \[ 
\begin{align*} 
K (Z =19) &= 1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 \\
Kr (Z = 36) &= 1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 \\
Cr (Z = 24) &= 1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^5 \\
Fe^{2+} (Z = 26) &= 1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^5 \\
Cl^- (Z = 17) &= 1s^2 2s^2 2p^6 \\
Si (Z = 14) &= 1s^2 2s^2 2p^6 \\
\end{align*} 
\]

Note: Valence shell electron configuration is in bold.

(b) Valence shell electron configurations are given below.

\[
\begin{align*} 
Cs^+ (Z = 55) &= [Kr]5s^2 4d^{10} 5p^6 \\
Ir^{3+} (Z = 77) &= [Xe]4f^{14} 5d^6 \\
Sr (Z = 38) &= [Kr]5s^2 \\
Mn^{2+} (Z = 25) &= [Ar]3d^5 \\
S^{2-} (Z = 16) &= [Ar]3s^2 3p^6 \\
\end{align*} 
\]

4. The atomic number and the number of shells increase from Be to Ba. Therefore the size of the atoms increases from Be to Ba.

5. The atomic radii generally decrease as we move from Na to Cl. As atomic number increases, the effective nuclear charge (positive) increases. Therefore, the attraction of electrons to the nucleus is stronger when moving from left to right.

6. \[ 
\begin{align*} 
Li &= 1s^2 2s^1, & \text{Li}^+ &= 1s^2 \\
\end{align*} 
\]

Once lithium forms a cation by giving out the outermost electron, the number of shells is reduced to 1. The effective nuclear charge has increased. As a result, the two 1s electrons are more strongly attracted to the nucleus causing a significant reduction in its radius.

7. After receiving an electron, iodine atom generates the negatively charged iodide ion ($I^-$). This incoming electron enters the outermost sub-shell (5p) which is now completely filled. The electron cloud is expanded due to electron-electron repulsion. Therefore, the radius of the $I^-$ ion is much larger than that of the iodine atom.

8. a) Metallic Character: N < P < As < Sb;  
b) Non-metallic Character: Sr < Ca < Mg < Be
Mg = 1s^2 2s^2 2p^6 3s^2  
Al = 1s^2 2s^2 2p^6 3s^2 3p^1

The first ionization energy means the energy required to remove an electron from the gaseous atom. Mg has two outermost electrons in the completely filled s sub-shell. Since the electron configurations with filled shells/sub-shells show an extra stability than those with unfilled shells/sub-shells, the removal of an electron from Mg is not easy. In the case of Al, the electron is removed from the unfilled p sub-shell and the removal of an electron leads to a filled electron configuration, therefore it is much easier to remove an electron from Al. Therefore, the first ionization energy of Mg is greater than that of Al.

Cl = 1s^2 2s^2 2p^6 3s^2 3p^5  
Ar = 1s^2 2s^2 2p^6 3s^2 3p^6

By getting one electron, Cl atom can easily attain the stable noble gas configuration, the electron affinity of Cl atom is a negative value (energy is liberated when an electron is added). But it is difficult to give an electron to Ar, since it destroys the stable noble gas configuration. Therefore energy should be supplied from outside to force an electron to combine with Ar(g). Hence electron affinity of Ar is positive.

Electronegativity:  
(a) Li < H < C < N < O < F  
(b) Ba < Ca < Mg < Be  
(c) I < Br < Cl < F

<table>
<thead>
<tr>
<th>Group 2</th>
<th>m. pt./°C</th>
<th>Group 16</th>
<th>m. pt./°C</th>
</tr>
</thead>
<tbody>
<tr>
<td>Be</td>
<td>1283</td>
<td>O</td>
<td>−219</td>
</tr>
<tr>
<td>Mg</td>
<td>650</td>
<td>S</td>
<td>114</td>
</tr>
<tr>
<td>Ca</td>
<td>850</td>
<td>Se</td>
<td>217</td>
</tr>
<tr>
<td>Sr</td>
<td>770</td>
<td>Te</td>
<td>450</td>
</tr>
<tr>
<td>Ba</td>
<td>710</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

In general, the melting points of the Group 2 elements decrease with the increase of atomic number. The atoms are held together by weak metallic bonds and the strength of bonding decreases when moving down the group. Be atoms are packed more compactly (since its metallic bonds are stronger than others) and the melting point of Be is significantly high. The melting point of Mg is unexpectedly lower than Ca.

In Group 16, the oxygen has the lowest melting point. In general, melting points increase as we go down the group because the covalent/ionic bond strengths increase with increasing atomic number. Since the strength of the forces which hold the atoms in the solid state is high, it is difficult to form liquid by breaking these bonds.

(a) **Group 1 elements**

For example, the electron configuration of Na = 1s^2 2s^2 2p^6 3s^1

Na attains its stable electron configuration by giving away the outermost electron, hence the oxidation number of the Group 1 elements is always +1

(b) **Group 2 elements**

For example, the electron configuration of Mg = 1s^2 2s^2 2p^6 3s^2

By giving away 2 electrons it can attain the stable noble gas configuration. Therefore, the oxidation number of the Group 2 elements is always +2.
(c) **Group 16 elements**
For example, the electron configuration of O = 1s$^2$ 2s$^2$ 2p$^4$
Oxidation number of O in its compound is often $-2$, because it tries to attain noble gas configuration by gaining 2 electrons.

S = 1s$^2$ 2s$^2$ 2p$^6$ 3s$^2$ 3p$^4$
Oxidation number of S varies from $-2$ to $+6$; most common oxidation numbers are $-2$, $+4$ and $+6$. This is due to the availability of $d$ orbitals.

14 The electronegativity of an atom is a measure of the tendency of the atom to attract electrons when it is covalently bonded to other atom(s).

15 a) He (2372) and Li (520)
The first ionization energy of He is larger than that of Li. The reason is, He has two outermost electrons in a stable filled $s$ sub shell. But Li has the outermost electron in a half filled $s$ orbital. The removal of one electron from the filled $s$ sub shell is much harder than removal of a single electron from the half-filled sub shell.

b) Li (520) and Be (900)
Li has one outermost electron in the $s$ sub shell while Be has two outermost electrons in $s$ sub shell. Since the removal of one electron from the filled $s$ sub shell is much harder than the removal of a single electron from a half filled sub shell, Be has larger first ionization energy than that of Li.

c) Be (900) and B (800) ; Be = 1s$^2$ 2s$^2$, B = 1s$^2$ 2s$^2$ 2p$^1$
The first ionization energy of Be is larger than that of B. The reason is that Be has two outermost electrons in stable filled $s$ sub shell. But the outermost electron of B is in the $p$ sub shell which is not filled. Therefore the removal of one electron from stable $s$ sub shell is difficult and hence the first ionization energy of Be is larger than that of B.

16 N (Z = 7) = 1s$^2$2s$^2$ 2p$^3$
The $5^{th}$ ionization energy of N is the energy required to remove an electron from the tetravalent gaseous cation. These electrons are removed from the $s$ sub shell which has one electron (half-filled sub-shell). The $6^{th}$ ionization energy of N means the energy required for removing an electron from the gaseous pentavalent cation and these electrons have to be removed from the most stable filled $s$ sub shell. Removal of an electron from the filled sub shell is much harder because this 1s sub shell is situated very much closer to the nucleus. Therefore, the $6^{th}$ ionization energy of N is significantly high.

17 The amount of energy released when an electron is added to a neutral gaseous atom.

18 The electron affinity increases from left to right of the Periodic Table as the effective nuclear charge increases and the size of the atom decreases, from left to right.

19 Cr$_2$O$_3$, CrO$_2$ and CrO$_3$

20 VCl$_2$, VCl$_3$ and VCl$_4$. 

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Study Questions

1. Which element has a single electron in its outermost sub-shell?
   1) Ca  2) C  3) P  4) Ge  5) Cu

2. The element with the electron configuration of s\(^2\)p\(^6\)d\(^{10}\)s\(^1\) is
   1) Mo  2) K  3) Fe  4) Ag  5) Zn

3. Number of d electrons in the Mn\(^{2+}\) ion is
   1) 3  2) 4  3) 5  4) 6  5) 7

4. The ion/element which is isoelectronic with the Fe\(^{3+}\) ion is
   1) Cr\(^{3+}\)  2) Cr  3) Mn\(^{2+}\)  4) Mn  5) Co\(^{2+}\)

5. The highest oxidation state of Mn is
   1) +2  2) +4  3) +5  4) +6  5) +7

6. The highest oxidation state of fluorine is
   1) -1  2) 0  3) +1  4) +5  5) +7

7. The highest oxidation state of chlorine is
   1) -1  2) 0  3) +1  4) +5  5) +7

8. The molecular formula of the oxide formed by the highest oxidation state of the element X with the atomic number 17 is:
   1) X\(_2\)O  2) XO\(_2\)  3) XO\(_3\)  4) X\(_2\)O\(_5\)  5) X\(_2\)O\(_7\)

9. Which one of the following is the smallest in size?
   1) N\(^{3-}\)  2) O\(^{2-}\)  3) F\(^-\)  4) Na  5) K

10. Which one of the following electron configurations belongs to an alkaline earth element?
    1) 1s\(^2\)  2) ns\(^2\)np\(^l\)  3) ns\(^2\)np\(^6\)  4) nd\(^{10}\)(n+1)s\(^2\)  5) None of the above

11. The element/ion isoelectronic with Cr\(^{3+}\) is
    1) V\(^{2+}\)  2) Ti  3) Mn\(^{2+}\)  4) Fe\(^{3+}\)  5) B\(^{2+}\)

12. Which one of the following atoms has the d\(^7\)s\(^2\) electron configuration?
    1) Fe  2) Co  3) Ni  4) Cr  5) Zn

13. The ion which does not form a coloured compound is
    1) Zn\(^{2+}\)  2) Cr\(^{2+}\)  3) Mn\(^{2+}\)  4) Fe\(^{2+}\)  5) Ni\(^{2+}\)

14. Which one of the following elements is a metalloid?
    1) Sb  2) Sr  3) Se  4) Ga  5) Bi

15. Which one of the following elements is a non-metal?
    1) Se  2) As  3) Pt  4) Ga  5) B
16. Which element has the lowest first ionization energy?
1) N  2) O  3) F  4) Ni  5) He

17. The 1\textsuperscript{st}, 2\textsuperscript{nd}, 3\textsuperscript{rd}... ionization energies of an element are 685, 1320, 2240, 8950, 11600, 14500 kJ mol\textsuperscript{-1} respectively. The Group number of this element is
1) 3  2) 13  3) 14  4) 15  5) 16

18. The molecule which exhibits hydrogen bonding is
1) CH\textsubscript{3}I  2) N\textsubscript{2}  3) H\textsubscript{2}  4) HF  5) CH\textsubscript{4}

19. The removal of one electron is easiest in
1) Ni  2) Na\textsuperscript{+}  3) Mg\textsuperscript{2+}  4) Al\textsuperscript{3+}  5) O\textsuperscript{2-}

20. Which one of the following ions has the smallest ionic radius?
1) Al\textsuperscript{3+}  2) Ga\textsuperscript{3+}  3) Se\textsuperscript{2-}  4) N\textsuperscript{3-}  5) O\textsuperscript{2-}

21. The lowest oxidation state of the element X is -3. The valence electron configuration of X is
1) s\textsuperscript{2}p\textsuperscript{1}  2) s\textsuperscript{2}p\textsuperscript{3}  3) d\textsuperscript{2}s\textsuperscript{1}  4) d\textsuperscript{1}s\textsuperscript{2}  5) s\textsuperscript{2}p\textsuperscript{5}

22. The highest oxidation state of the element M is +3. The valence electron configuration of M is,
1) s\textsuperscript{2}p\textsuperscript{3}  2) d\textsuperscript{3}s\textsuperscript{2}  3) d\textsuperscript{2}s\textsuperscript{0}  4) s\textsuperscript{1}p\textsuperscript{2}  5) d\textsuperscript{1}s\textsuperscript{2}

23. Select the element with the highest second ionization energy.
1) Na  2) Mg  3) Ca  4) Be  5) Ba

24. Which element has a single electron in its d sub-shell?
1) Sc  2) Ti  3) V  4) Cr  5) Mn

25. Which one is not involved in the development of the Periodic Table?
1) Newlands  2) Mendeleev  3) Rutherford  4) Gay-Lussac  5) Mayer

26. What is the Group number of the element with the valence electron configuration ns\textsuperscript{2}np\textsuperscript{5}?
1) Group 1  2) Group 5  3) Group 7  4) Group 13  5) Group 17

27. The Group number of the element with the atomic number 33 is
1) 5  2) 14  3) 15  4) 16  5) 17

28. The concept of triads was introduced by
1) Dobereiner  2) Newlands  3) Rutherford  4) Mayer  5) Bohr

29. Which element (M) forms the oxide with the molecular formula MO\textsubscript{3}?
1) Be  2) Al  3) Se  4) Si  5) C

30. “Octet Rule” was introduced by
1) Newlands  2) Mayer  3) Bohr  4) Dalton  5) Rutherford

31. Which one of the following set is a Dobereiner’s triad?
1) Na, K, Cs  2) Cl, Br, I  3) Mg, Ca, Sr  4) S, Se, Te  5) Be, B, C
32. The outermost electron configuration of halogens is
   1) (n-1)d^{10}ns^{2}np^{5}  2) ns^{2}np^{6}  3) (n-1)d^{10}ns^{2}np^{5}
   4) ns^{2}np^{4}  5) None of the above

33. The atomic number of the element which has a single electron in its outermost shell is
   1) 5  2) 14  3) 20  4) 25  5) 29

34. The correct order of increasing the ionic radii is
   1) Na<sup>+</sup> < Mg<sup>2+</sup> < Ca<sup>2+</sup>  2) Mg<sup>2+</sup> < Na<sup>+</sup> < Ca<sup>2+</sup>
   3) Ca<sup>2+</sup> < Mg<sup>2+</sup> < Na<sup>+</sup>  4) Mg<sup>2+</sup> < Ca<sup>2+</sup> < Na<sup>+</sup>
   5) Ca<sup>2+</sup> < Na<sup>+</sup> < Mg<sup>2+</sup>

35. The element which is more likely to lose two electrons easily is
   1) Cs  2) Al  3) Ba  4) Be  5) K

36. Which element has the lowest first ionization energy?
   1) K  2) Mg  3) Ca  4) Rb  5) Sr

37. The element with the highest second ionization energy is
   1) Na  2) Mg  3) Ca  4) Be  5) Li

38. X, Y and Z are consecutive elements with increasing atomic numbers. Element X is a noble gas. The symbol of the ion of element Z in its compounds is
   1) Z<sup>−</sup>  2) Z<sup>2−</sup>  3) Z<sup>3−</sup>  4) Z<sup>2+</sup>  5) Z<sup>3+</sup>

39. Which one of the following has the electron configuration of N<sup>3−</sup>?
   1) Cl<sup>−</sup>  2) Li<sup>+</sup>  3) Ar  4) Al<sup>3+</sup>  5) Be<sup>2+</sup>

40. Which pair of elements shows diagonal relationship?
   1) B, Al  2) Be, Mg  3) Li, Mg  4) C, Si  5) Be, B

41. Consider the following ions.
   a) Cl<sup>−</sup>  b) N<sup>3−</sup>  c) Al<sup>3+</sup>  d) Be<sup>2+</sup>
   The ions with the electronic configuration of Ne are
   1) a & b only  2) b & c only  3) c & d only  4) a & d only  5) a, b & c

42. Consider the following pairs of elements.
   a) Be, Mg  b) Li, Be  c) Be, Al  d) Li, Mg
   The pairs of elements that show the diagonal relationship are
   1) a & b  2) b & c  3) c & d  4) a & d  5) a, b & c

43. Which one of the following elements has the highest first ionization energy?
   1) Si  2) P  3) As  4) Ge  5) C

44. Among the following elements, select the largest atom.
   1) Sr  2) Ba  3) Cs  4) Rb  5) Ge

45. The valencies of an element with an outer electron configuration ns<sup>2</sup>np<sup>4</sup> are
   1) 1 and 4  2) 2 and 1  3) 2 and 5  4) 2 and 6  5) 1 and 5
Glossary

**Alkali metal**: The Group 1 elements, lithium (Li), sodium (Na), potassium (K), rubidium (Rb), cesium (Cs), and francium (Fr).

**Alkaline earth metal**: The Group 2 elements, beryllium (Be), magnesium (Mg), calcium (Ca), strontium (Sr), barium (Ba), and radium (Ra).

**Atom**: The smallest object that retains properties of an element. It is composed of electrons and a nucleus containing protons and neutrons.

**Block**: A region of the periodic table that corresponds to the type of subshell (s, p, d, or f) being filled during the Aufbau construction of electron configurations.

**Electron affinity**: The amount of energy released when an electron is added to a neutral gaseous atom.

**Electronegativity**: The tendency of an atom to attract electrons to itself when combined in a molecule.

**Element**: A substance consisting of only one type of atoms.
First ionization energy: The energy needed to remove an electron from an isolated, neutral gaseous atom.

First ionization energy: The energy needed to remove an electron from an isolated, neutral gaseous atom.

Group: A vertical column on the periodic table.

Halogen: Elements of Group 18. Fluorine (F), chlorine (Cl), bromine (Br), iodine (I), and astatine (At) are known at this time.

Ionic radius: The radii of anions and cations in crystalline ionic compounds, as determined by consistently partitioning the center-to-center distance of ions in those compounds.

Ionisation energy: The energy needed to remove an electron from a gaseous atom or ion.

Liquefaction: The act or process of turning a gas into a liquid.

Mass number: The total number of protons and neutrons in the nucleus of an atom.

Metal: It is a substance that conducts heat and electricity, it is shiny and can be hammered into sheets or drawn into wire. Metals lose electrons easily to form cations.
Metalloid: It is a substance that exhibits properties in between to metals and non-metals.

Nitrogen Fixation: The process by which free nitrogen from the air is combined with other elements to form inorganic compounds.

Nonmetal: It is a substance that conducts heat and electricity poorly, is brittle or waxy or gaseous, and cannot be hammered into sheets or drawn into wire. Nonmetals gain electrons easily to form anions.

Neutron: A neutral particle in the nucleus of an atom.

Orbital: A region around the nucleus where there is a high probability of finding an electron.

Oxidation number: A number assigned to each atom to help keep track of the electrons during a redox-reaction.

Paramagnetic: A substance that shows magnetic properties when placed in a magnetic field.
Periodic Table : An arrangement of the elements according to increasing atomic number that shows relationships between element properties.

Periodic trend : A regular variation in element properties with increasing atomic number that is ultimately due to regular variations in atomic structure.

Period : Horizontal rows in the periodic table.

Proton : A positively charged particle in the nucleus of an atom.

Reducing agent : A substance that causes another substance to undergo reduction and that is oxidized in the process.

Second ionization energy: The energy needed to remove an electron from an isolated singly charged positive gaseous ion.

Subshell : One part of a energy level, each of which can hold different numbers of electrons.

Transition metal : An element with an incomplete d subshell.

Valence electrons : The electrons in the outermost shell of an atom.
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Images

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16. Potassium:
https://upload.wikimedia.org/wikipedia/commons/thumb/a/a4/Potassium-2.jpg/220px-Potassium-2.jpg

17. Rubidium:
https://upload.wikimedia.org/wikipedia/commons/thumb/c/c9/Rb5.JPG/220px-Rb5.JPG

18. Cesium:
https://upload.wikimedia.org/wikipedia/commons/thumb/3/3d/Cesium.jpg/220px-Cesium.jpg

19. Francium:

20. Beryllium:
https://upload.wikimedia.org/wikipedia/commons/thumb/0/0c/Be-140g.jpg/150px-Be-140g.jpg

21. Magnesium:

22. Calcium:
https://upload.wikimedia.org/wikipedia/commons/thumb/9/96/Calcium_unter_Argon_Schutzgasatmosph%C3%A4re.jpg/220px-Calcium_unter_Argon_Schutzgasatmosph%C3%A4re.jpg

23. Strontium:

24. Barium:
https://upload.wikimedia.org/wikipedia/commons/thumb/1/16/Barium_unter_Argon_Schutzgas_Atmosph%C3%A4re.jpg/220px-Barium_unter_Argon_Schutzgas_Atmosph%C3%A4re.jpg

25. Radium:

26. Boron:
https://upload.wikimedia.org/wikipedia/commons/thumb/1/19/Boron_R105.jpg/220px-Boron_R105.jpg

27. Aluminium:
https://upload.wikimedia.org/wikipedia/commons/thumb/5/5d/Aluminium-4.jpg/220px-Aluminium-4.jpg
29. Indium: 
https://encrypted-tbn3.gstatic.com/images?q=tbn:ANd9GcTDG7j_sA9nmUtwItcpDSk8SEI36n-0lEjxtBUPMNM3X8vGoEJs

30. Thallium: 

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   i. https://encrypted-tbn1.gstatic.com/images?q=tbn:ANd9GcTAW_n9_FWk9sMrxE1rm4IghWHfsRfjnn-tReWuTFOqUsqEJCp
   ii. https://encrypted-tbn3.gstatic.com/images?q=tbn:ANd9GcSz1CY11Y7I5sYtXUhGRVtzGKCMaHEAlqU9opmJfJ0RkTo_vvUk
   iii. http://www.docbrown.info/ks3chemistry/gifs/bucky.gif

32. Carbon: 

33. Silicon: 
https://upload.wikimedia.org/wikipedia/commons/thumb/e/e9/SiliconCroda.jpg/220px-SiliconCroda.jpg

34. Germanium: 
https://upload.wikimedia.org/wikipedia/commons/thumb/0/08/Polycrystalline-germanium.jpg/220px-Polycrystalline-germanium.jpg

35. Tin: 
https://upload.wikimedia.org/wikipedia/commons/thumb/2/2b/Sn-Alph-Beta.jpg/220px-Sn-Alpha-Beta.jpg

36. Lead: 
https://upload.wikimedia.org/wikipedia/commons/thumb/e/e6/Lead_electrolytic_and_1cm3_cube.jpg/220px-Lead_electrolytic_and_1cm3_cube.jpg

37. Nitrogen: 
https://upload.wikimedia.org/wikipedia/commons/thumb/d/d2/Liquidnitrogen.jpg/220px-Liquidnitrogen.jpg

38. Phosphorus: 
https://upload.wikimedia.org/wikipedia/commons/thumb/8/88/PhosphComby.jpg/220px-PhosphComby.jpg

39. Arsenic: 
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41. Bismuth: https://upload.wikimedia.org/wikipedia/commons/thumb/e/ef/Bismuth_crystals_and_1cm3_cube.jpg/220px-Bismuth_crystals_and_1cm3_cube.jpg

42. Oxygen: https://upload.wikimedia.org/wikipedia/commons/1/1b/Liquid_Oxygen.png


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Course Team

Author
Prof. K. Sarath D. Perera
(Senior professor in Chemistry)

Content Editor
Dr. Sithy S. Iqbal

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