

## 1. Details of Module and its structure

Module Detail	
Subject Name	Chemistry
Course Name	Chemistry 01 (Class XI, Semester 01)
Module Name/Title	Classification of elements & periodicity in properties: Part 3
Module Id	kech010303
Pre-requisites	Periodic law, atomic number, electronic configuration & periodic classification
Objectives	After going through this module, the learners will be able to: <ol style="list-style-type: none"><li>1. Use appropriate scientific vocabulary to communicate ideas related to properties of atoms, namely - atomic and ionic radii, ionisation enthalpy, electronegativity and electron gain enthalpy.</li><li>2. Explain the relation between ionisation enthalpy and metallic character.</li><li>3. Recognise the periodic trends in the physical and chemical properties of the elements.</li><li>4. Compare the reactivity of elements and correlate it with their occurrence.</li></ol>
Keywords	Atomic Radii, Ionic Radii, Ionisation Enthalpy, Electron Gain Enthalpy, Electronegativity, Metallic & Non-Metallic Character, Oxidation States and Chemical Reactivity

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**1. Introduction:** Many trends in physical and chemical properties of elements are observed from top to bottom in a group or from left to right across a period in the Periodic Table.

In a period, chemical reactivity tends to be high in Group 1 metals, decreases in elements towards the middle of the table, and then increases to a maximum in the Group 17 non-metals. Similarly, within a group of alkali metals, the reactivity increases on moving down the group, whereas within a group of non-metals (say halogens), the reactivity decreases down the group. To explain these periodic trends, a study of atomic structure and properties of the atom is imperative. This module explains the periodic trends in certain physical and chemical properties in terms of number of electrons and energy levels of an element.

**2. Trends in Physical Properties:** Out of the many physical properties of elements such as melting and boiling points, heats of fusion and vaporization, enthalpy of atomization, etc. which show periodic variations, only the periodic trends with respect to atomic and ionic radii, ionisation enthalpy, electron gain enthalpy and electronegativity are discussed in this module.

### a) Atomic Radius

The size of an atom ( $\sim 1.2 \text{ \AA}$  i.e.,  $1.2 \times 10^{-10} \text{ m}$  in radius) is extremely small and the electron cloud surrounding the atom does not have a sharp boundary. Therefore the determination of the atomic radii is never accurate. However, to know the radii of an atom of a non-metallic element, the distance between two atoms when they are bound together by a single bond in a covalent molecule must be estimated and from this value, the “Covalent Radius” of the element can be calculated.

The bond distance in the chlorine molecule ( $\text{Cl}_2$ ) is 198 pm and half of this distance (99 pm), is taken as the atomic radius of chlorine.

For metals, the term “Metallic Radius” is taken as half of the inter-nuclear distance separating the metal cores in the metallic crystal.

The distance between two adjacent copper atoms in solid copper is 256 pm; hence the metallic radius of copper is assigned a value of 128 pm.

To avoid any confusion, the term, 'Atomic Radius' is referred to both covalent and metallic radius depending on whether the element is a non-metal or a metal. Atomic radii can be measured by X-ray or other spectroscopic methods. The atomic radii of a few elements are listed in Table 1 & Table 2. Two trends are possible. These trends can be explained in terms of nuclear charge and energy level. The atomic radii generally decreases across a period as illustrated in Fig.1 for the elements of the second period. It is because that within a period as the atomic number increases the electrons are being added in the same valence shell which results in an increase in effective nuclear charge and in turn the electrons are attracted by the nucleus to a greater extent. Thus shrinkage in atomic radii is observed.

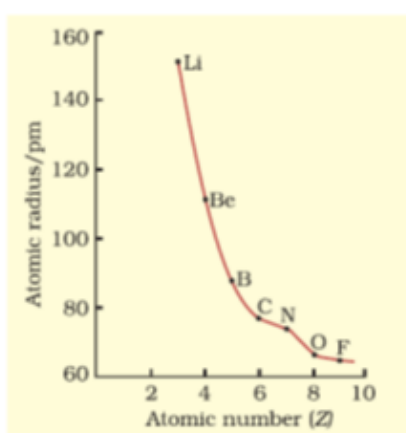


Fig 1: Variation of atomic radius with atomic number across the second period

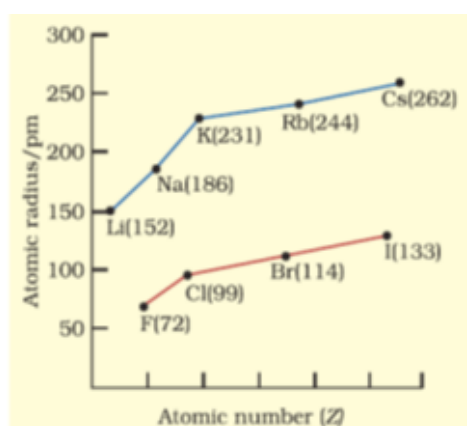


Fig. 2: Variation of atomic radius with atomic number for alkali metals and halogens

Within a group (vertical column) of the periodic table, the atomic radius increases regularly as the atomic number increases as illustrated in Fig. 2 (for alkali metals and halogens). From top to bottom in a group, the principal quantum number ( $n$ ) increases, the valence electrons are farther away from the nucleus as these enter a new shell with each descend. The inner energy levels are completely filled with electrons, which serve to shield the outer electrons from the pull of the nucleus. Thus, the size of the atom increases as reflected in the atomic radii.

[Note: the atomic radii of noble gases are not considered because being monoatomic; their (non-bonded radii) values are very large. In fact radii of noble gases should be compared with the van der Waals radii and not with the covalent radii of other elements].

Table 1: Atomic Radii/pm Across the Periods

Atom (Period II)	Li	Be	B	C	N	O	F
Atomic radius	152	111	88	77	74	66	64
Atom (Period III)	Na	Mg	Al	Si	P	S	Cl
Atomic radius	186	160	143	117	110	104	99

Table 2: Atomic Radii/pm down a Group

Atom (Group I)	Atomic Radius	Atom (Group 17)	Atomic Radius
Li	152	F	64
Na	186	Cl	99
K	231	Br	114
Rb	244	I	133
Cs	262	At	140

### b) Ionic Radius

The removal of an electron from an atom results in the formation of a cation, whereas a gain of an electron by an atom leads to the formation of an anion. The ionic radii can be estimated by measuring the distances between cations and anions in ionic crystals. In general, the ionic radii of elements exhibit the same trend as the atomic radii. A cation is smaller than its parent atom because it has fewer electrons while its nuclear charge remains the same. The size of an anion will be larger than that of the parent atom because the addition of one or more electrons would result in increased repulsion among the electrons and a decrease in effective nuclear charge, thus the outermost electrons are not attracted by the nucleus as much and hence an increase in the size is observed for an anion.

The ionic radius of fluoride ion ( $F^-$ ) is 136 pm, while the atomic radius of fluorine atom is only 64 pm. On the other hand, the ionic radius of sodium ion ( $Na^+$ ) is 95 pm, while the atomic radius of sodium atom is 186 pm.

**Isoelectronic species:** The atoms and ions which contain the same number of electrons are called as isoelectronic species.

$O^{2-}$ ,  $F^-$ ,  $Na^+$  and  $Mg^{2+}$  have 10 electrons each and are therefore are isoelectronic species. However, their radii would be different because these have different nuclear charges. Also the cation with the greater positive charge will have a smaller radius because of the greater attraction of the electrons to the nucleus. Anion with the greater negative charge will have the larger radius because the net repulsion of the electrons will exceed the nuclear charge and the anion will expand in size.

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**Problem 1:** Which of the following species will have the largest and the smallest radii?  
Mg, Mg<sup>2+</sup>, Al, Al<sup>3+</sup>

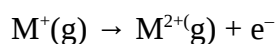
**Solution:** Atomic radius decreases across a period. Cations are smaller than their parent atoms. Among isoelectronic species, the one with the larger positive nuclear charge will have a smaller radius. Hence the largest species is Mg will have largest radii and Al<sup>3+</sup> will have smallest radii.

### (c) Ionisation Enthalpy

A quantitative measure of the tendency of an element to lose electron is given by its ionisation Enthalpy. It represents the energy required to remove an electron from an isolated gaseous atom in its ground state. In other words, the first ionisation enthalpy for an element M is the enthalpy change ( $\Delta_i H$ ) for the reaction depicted in equation given below.



The ionisation enthalpy is expressed in units of kJ mol<sup>-1</sup>. The second ionisation enthalpy can be defined as the energy required for removing the second most loosely bound electron; it is the energy required to carry out the reaction shown in equation below.



Energy is always required to remove electrons from an atom and hence ionisation enthalpies are always positive. The second ionisation enthalpy will be higher than the first ionisation enthalpy, because it is more difficult to remove an electron from a positively charged ion than from a neutral atom. In the same way the third ionisation enthalpy will be higher than the second and so on. The term “ionisation enthalpy”, unless specified, is taken as the first ionisation enthalpy. The first ionisation enthalpies of elements having atomic numbers up to 60 are plotted in Fig. 3. The periodicity of the graph is quite remarkable. You will find maxima at the noble gases which have closed electron shells and very stable electron configurations. On the other hand, minima occur at the alkali metals and their low ionisation enthalpies can be correlated with the high reactivity of alkali metals.

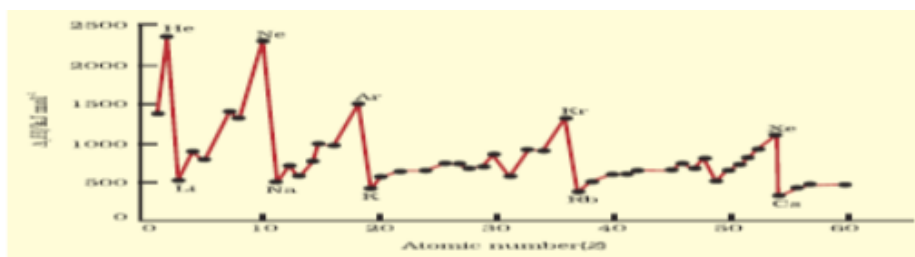


Fig. 3. Variation of first ionization enthalpies ( $\Delta_1H$ ) with atomic number for elements with  $Z = 1$  to 60

In the Fig. 3 above, two trends are noticed, the first ionisation enthalpy generally increases as we go across a period and decreases as we descend in a group. These trends are illustrated in Figs. 4(a) and 4(b) respectively for the elements of the second period and the first group of the periodic table. The ionisation enthalpy and atomic radius are closely related properties. These trends can be understood by considering two factors: (i) the attraction of electrons towards the nucleus, and (ii) the repulsion of electrons from each other. The effective nuclear charge experienced by a valence electron in an atom will be less than the actual charge on the nucleus because of “shielding” or “screening” of the valence electron from the nucleus by the intervening core electrons.

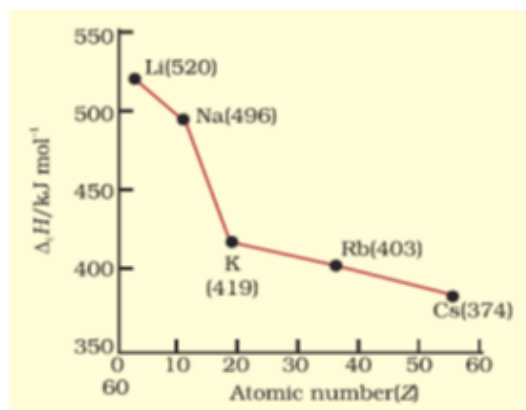
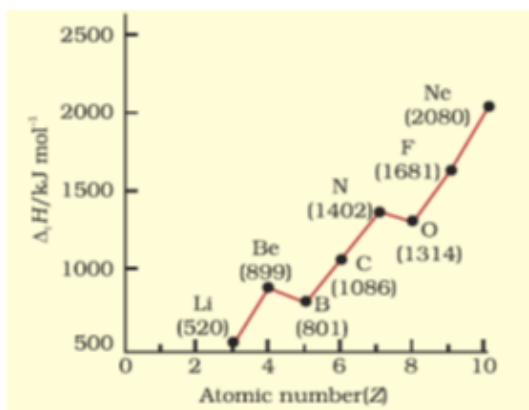


Fig. 4 (a) First ionization enthalpies ( $\Delta_1H$ ) of elements of the second period as a function of atomic number ( $Z$ ) and Fig. 4 (b)  $\Delta_1H$  of alkali metals as a function of  $Z$ .

For example, the 2s electron in lithium is shielded from the nucleus by the inner core of 1s electrons. As a result, the valence electron experiences a net positive charge which is less than the actual charge of +3. In general, shielding is effective when the orbitals in the inner shells are completely filled. This situation occurs in the case of alkali metals which have single outermost ns-electron preceded by a noble gas electronic configuration. When we move from lithium to fluorine across the second period, successive electrons are added to orbitals in the same principal quantum level and the shielding of the nuclear charge by the inner core of electrons does not increase very much to compensate for the increased attraction of the

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electron to the nucleus. Thus, across a period, increasing nuclear charge exceeds the shielding. Consequently, the outermost electrons are held more and more tightly and the ionisation enthalpy increases across a period. As we go down a group, the outermost electron being farther away from the nucleus, there is an increased shielding of the nuclear charge by the electrons in the inner levels. In this case, increase in shielding exceeds the increasing nuclear charge and the removal of the outermost electron requires less energy down a group. From Fig. 4(a), it is evident that the first ionisation enthalpy of boron ( $Z = 5$ ) is slightly less than that of beryllium ( $Z = 4$ ) even though the former has a greater nuclear charge. Considering the same principal quantum level, an s-electron is attracted to the nucleus more than a p-electron. In beryllium, the electron removed during the ionisation is an s-electron whereas the electron removed during ionisation of boron is a p-electron. The penetration of a 2s-electron to the nucleus is more than that of a 2p-electron; hence the 2p electron of boron is more shielded from the nucleus by the inner core of electrons than the 2s electrons of beryllium. Therefore, it is easier to remove the 2p-electron from boron compared to the removal of a 2s- electron from beryllium. Thus, boron has a smaller first ionisation enthalpy than beryllium. Another “anomaly” is the smaller first ionisation enthalpy of oxygen compared to nitrogen. This is because in the nitrogen atom, three 2p-electrons are in different atomic orbitals (Hund’s rule) where each orbital of 2p is half filled while in the oxygen atom, two of the four 2p-electrons occupy the same 2p-orbital resulting in increased electron-electron repulsion. Consequently, it is easier to remove the fourth 2p-electron from oxygen than it is, to remove one of the three 2p-electrons from nitrogen.

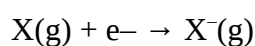
**Problem 2:** The first ionisation enthalpy ( $\Delta_i H$ ) values of the elements in third period Na, Mg and Si are respectively 496, 737 and 786  $\text{kJ mol}^{-1}$ . Predict and justify whether the first  $\Delta_i H$  value for Al will be more close to 575 or 760  $\text{kJ mol}^{-1}$ ?

**Solution:** The first ionisation enthalpy for Al will be close to 575  $\text{kJ mol}^{-1}$  and the value of  $\Delta_i H$  for Al should be lower than that of Mg because of effective shielding of 3p-electrons from the nucleus by 3s-electrons.

#### (d) Electron Gain Enthalpy

When an electron is added to a neutral gaseous atom (say X) to convert it into a negative ion, the enthalpy change accompanying the process is defined as the Electron Gain Enthalpy

$\Delta_{eg} H$  Electron gain enthalpy provides a measure of the ease with which an atom adds an electron to form anion as represented by the equation:



Depending on the element, the process of adding an electron to the atom can be either endothermic or exothermic. For many elements energy is released when an electron is added to the atom and the electron gain enthalpy is negative. For example, group 17 elements (the halogens) have very high negative electron gain enthalpies because they can attain stable noble gas electronic configurations by picking up an electron. On the other hand, noble gases have large positive electron gain enthalpies because the electron has to enter the next higher principal quantum level leading to a very unstable electronic configuration. It may be noted that electron gain enthalpies have large negative values toward the upper right of the periodic table preceding the noble gases. The variation in electron gain enthalpies of elements is less systematic than for ionisation enthalpies. As a general rule, electron gain enthalpy becomes more negative with increase in the atomic number across a period. The effective nuclear charge increases from left to right across a period and consequently it will be easier to add an electron to a smaller atom since the added electron on an average would be closer to the positively charged nucleus. We should also expect electron gain enthalpy to become less negative as we go down a group because the size of the atom increases and the added electron would be farther from the nucleus. This is generally the case (Table 3). However, electron gain enthalpy of O or F is less negative than that of the succeeding element. This is because when an electron is added to O or F, the added electron goes to the smaller  $n = 2$  quantum level and suffers significant repulsion from the other electrons present in this level. For the  $n = 3$  quantum level (S or Cl), the added electron occupies a larger region of space and the electron-electron repulsion is much less.

**Table 3. Electron Gain Enthalpies\* / (kJ mol<sup>-1</sup>) of Some Main Group Elements**

Group 1	$\Delta_{eg}H$	Group 16	$\Delta_{eg}H$	Group 17	$\Delta_{eg}H$	Group 0	$\Delta_{eg}H$
H	- 73					He	+ 48
Li	- 60	O	- 141	F	- 328	Ne	+ 116
Na	- 53	S	- 200	Cl	- 349	Ar	+ 96
K	- 48	Se	- 195	Br	- 325	Kr	+ 96
Rb	- 47	Te	- 190	I	- 295	Xe	+ 77
Cs	- 46	Po	- 174	At	- 270	Rn	+ 68

**Problem 3.** Which of the following will have the most negative and which the least negative electron gain enthalpy? P, S, Cl, F. Explain your answer.

**Solution:** Electron gain enthalpy generally becomes more negative across a period as we move from left to right. Within a group, electron gain enthalpy becomes less negative down a



group. However, adding an electron to the 2p-orbital leads to greater repulsion than adding an electron to the larger 3p-orbital. Hence the element with most negative electron gain enthalpy is chlorine; the one with the least negative electron gain enthalpy is phosphorus.

### (e) Electronegativity

A qualitative measure of the ability of an atom in a chemical compound to attract the shared pair of electrons towards itself is called electronegativity. Unlike ionisation enthalpy and electron gain enthalpy, it is not a measurable quantity. A number of numerical scales of electronegativity of elements viz., Pauling scale, Mulliken-Jaffe scale, Allred-Rochow scale have been developed and the Pauling scale is the most widely used. Linus Pauling, an American scientist, in 1922 assigned arbitrarily a value of 4.0 to fluorine, the element considered to have the greatest ability to attract electrons. Approximate values for the electronegativity of a few elements are given in Table 4 (a) & (b).

**Table 4 (a): Electronegativity Values (on Pauling scale) Across the Periods**

Atom (Period II)	Li	Be	B	C	N	O	F
Electronegativity	1.0	1.5	2.0	2.5	3.0	3.5	4.0
Atom (Period III)	Na	Mg	Al	Si	P	S	Cl
Electronegativity	0.9	1.2	1.5	1.8	2.1	2.5	3.0

**Table 4 (b): Electronegativity Values (on Pauling scale) Down a Family**

Atom (Group I)	Electronegativity Value	Atom (Group 17)	Electronegativity Value
Li	1.0	F	4.0
Na	0.9	Cl	3.0
K	0.8	Br	2.8
Rb	0.8	I	2.5
Cs	0.7	At	2.2

The electronegativity of any given element is not constant; it varies depending on the element to which it is bound and one could predict the nature of force that holds a pair of atoms together from the electronegativity values. Electronegativity generally increases across a period from left to right (say from lithium to fluorine) and decrease down a group (say from fluorine to astatine) in the periodic table. The attraction between the outer (or valence) electrons and the nucleus increases as the atomic radius decreases in a period. The electronegativity also increases. Thus the electronegativity values decrease with the increase in atomic radii down a group. The trend is similar to that of ionisation enthalpy. The increase

in electronegativities across a period is accompanied by an increase in non-metallic properties (or decrease in metallic properties) of elements. Similarly, the decrease in electronegativity down a group is accompanied by a decrease in non-metallic properties (or increase in metallic properties) of elements. All these periodic trends are summarised in Fig 5.

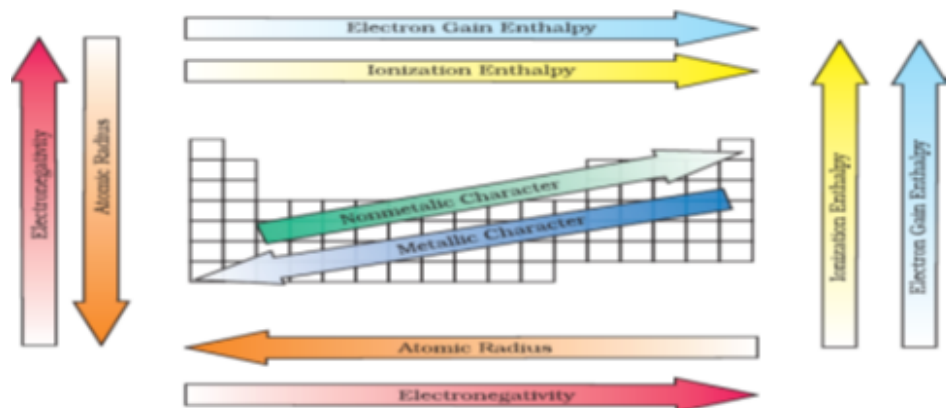


FIG. 5 The periodic trends of elements in the periodic table

**3. Periodic Trends in Chemical Properties:** Most of the trends in chemical properties of elements, such as diagonal relationships, inert pair effect, effects of lanthanoide contraction etc. will be dealt with along the discussion of each group in later units. In this section we shall study the periodicity of the valence state shown by elements and the anomalous properties of the second period elements (from lithium to fluorine).

**a) Periodicity of Valence or Oxidation States:**

The valency is the most characteristic property of the elements and can be understood in terms of their electronic configurations. The valency of elements is usually (though not necessarily) equal to the number of electrons in the outermost orbitals and / or equal to eight minus the number of outermost electrons. Nowadays, the term oxidation state is frequently used in place of valency. Consider the two oxygen containing compounds:  $\text{OF}_2$  and  $\text{Na}_2\text{O}$ . The order of electronegativity of the three elements in these compounds is  $\text{F} > \text{O} > \text{Na}$ . Each of the atoms of fluorine, with outer configuration  $2s^2 2p^5$ , shares one electron with oxygen in the  $\text{OF}_2$  molecule. Being highest electronegative element, fluorine is given oxidation state  $-1$ . Since there are two fluorine atoms in this molecule, oxygen with outer electronic configuration  $2s^2 2p^4$  shares two electrons with fluorine atoms and thereby exhibits oxidation state  $+2$ . In  $\text{Na}_2\text{O}$ , oxygen being more electronegative accepts two electrons, one from each of the two sodium atoms and, thus, shows oxidation state  $-2$ . On the other hand sodium with electronic configuration  $3s^1$  loses one electron to oxygen and is given oxidation state  $+1$ . Thus, the oxidation state of an element in a particular compound can be defined as the **charge**

acquired by its atom on the basis of electronegative consideration from other atoms in the molecule.

**Problem 4:** Using the Periodic Table, predict the formulas of compounds which might be formed by the following pairs of elements; (i) silicon and bromine (ii) aluminium and sulphur.

**Solution:** (i) Silicon is element of group 14 with a valency of 4; bromine belongs to the halogen family (Group 17) with a valency of 1. Hence the formula of the compound formed would be  $\text{SiBr}_4$ .

(ii) Aluminium belongs to group 13 with a valency of 3; sulphur belongs to group 16 elements with a valency of 2. Hence, the formula of the compound formed by these two elements would be  $\text{Al}_2\text{S}_3$ .

Some periodic trends observed in the valency of elements (hydrides and oxides) are shown in Table 5. There are many elements which exhibit variable valency. This is particularly characteristic of transition elements and actinoids.

**Table 5: Periodic Trends in Valence of Elements as shown by the Formulas of Their Compounds**

Group	1	2	13	14	15	16	17
Formula of hydride	LiH		$\text{B}_2\text{H}_6$	$\text{CH}_4$	$\text{NH}_3$	$\text{H}_2\text{O}$	$\text{HF}$
	NaH	$\text{CaH}_2$	$\text{AlH}_3$	$\text{SiH}_4$	$\text{PH}_3$	$\text{H}_2\text{S}$	$\text{HCl}$
	KH			$\text{GeH}_4$	$\text{AsH}_3$	$\text{H}_2\text{Se}$	$\text{HBr}$
				$\text{SnH}_4$	$\text{SbH}_3$	$\text{H}_2\text{Te}$	$\text{HI}$
Formula of oxide	$\text{Li}_2\text{O}$	$\text{MgO}$	$\text{B}_2\text{O}_3$	$\text{CO}_2$	$\text{N}_2\text{O}_3, \text{N}_2\text{O}_5$		-
	$\text{Na}_2\text{O}$	$\text{CaO}$	$\text{Al}_2\text{O}_3$	$\text{SiO}_2$	$\text{P}_4\text{O}_6, \text{P}_4\text{O}_{10}$	$\text{SO}_2$	$\text{Cl}_2\text{O}_7$
	$\text{K}_2\text{O}$	$\text{SrO}$	$\text{Ga}_2\text{O}_3$	$\text{GeO}_2$	$\text{As}_2\text{O}_3, \text{As}_2\text{O}_5$	$\text{SeO}_2$	-
		$\text{BaO}$	$\text{In}_2\text{O}_3$	$\text{SnO}_2$	$\text{Sb}_2\text{O}_3, \text{Sb}_2\text{O}_5$	$\text{TeO}_2$	-
				$\text{PbO}_2$	$\text{Bi}_2\text{O}_3$ -	-	-

#### b) Anomalous Properties of Second Period Elements:

Property	Element		
Metallic radius M/ pm	Li	Be	B
	152	111	88
	Na	Mg	Al
	186	160	143
Ionic radius $\text{M}^+$ / pm	Li	Be	
	76	31	
	Na	Mg	
	102	72	

The first element of each of the groups 1(lithium) and 2 (beryllium) and groups 13-17 (boron to fluorine) differs in many respects from the other members of their respective group. For example, lithium unlike other alkali metals, and beryllium unlike other alkaline earth metals,

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form compounds with pronounced covalent character; the other members of these groups predominantly form ionic compounds. In fact the behaviour of lithium and beryllium is more similar with the second element of the following group i.e., magnesium and aluminium, respectively. This sort of similarity is commonly referred to as **diagonal relationship** in the periodic properties. The anomalous behaviour of the first member of a group of elements in the s- and p-blocks compared to that of the subsequent members in the same group is attributed to their small size, large charge/ radius ratio and high electronegativity of the elements. In addition, the first member of group has only four valence orbitals (2s and 2p) available for bonding, whereas the second member of the groups have nine valence orbitals (3s, 3p, 3d). As a consequence of this, the maximum covalence of the first member of each group is 4 (e.g., boron can only form  $[\text{BF}_4]^-$  whereas the other members of the groups can expand their valence shell to accommodate more than four pairs of electrons e.g., aluminium forms  $[\text{AlF}_6]^{3-}$ ). Furthermore, the first member of p-block elements displays greater ability to form  $p\pi - p\pi$  multiple bonds to itself (e.g.,  $\text{C} = \text{C}$ ,  $\text{C} \equiv \text{C}$ ,  $\text{N} = \text{N}$ ,  $\text{N} \equiv \text{N}$ ) and to other second period elements (e.g.,  $\text{C} = \text{O}$ ,  $\text{C} = \text{N}$ ,  $\text{C} \equiv \text{N}$ ,  $\text{N} = \text{O}$ ) compared to subsequent members of the same group.

**Problem 5:** Are the oxidation state and covalency of Al in  $[\text{AlCl}(\text{H}_2\text{O})_5]^{2+}$  same ?

**Solution:** No. The oxidation state of Al is +3 and the covalency is 6.

### (c) Periodic Trends and Chemical Reactivity

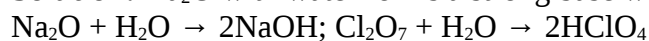
The periodic trends in certain fundamental properties such as atomic and ionic radii, electron gain enthalpy, ionisation enthalpy and valency are related to electronic configuration which is responsible for the chemical and physical properties of the elements. Thus the fundamental properties and the chemical reactivity of these elements are related to each other. The atomic and ionic radii generally decrease in a period from left to right, due to which the ionisation enthalpies increases (with some exceptions) and electron gain enthalpies become more negative across a period. Thus, the ionisation enthalpy of the extreme left element in a period is the least and the electron gain enthalpy of the element on the extreme right is the highest negative, which results into high chemical reactivity at the two extremes and the lowest in the centre (note: Noble gases having completely filled shells therefore, have positive electron gain enthalpy values). Thus, the maximum chemical reactivity at the extreme left (among alkali metals) is observed as a loss of an electron leading to the formation of a cation and at the extreme right (among halogens) is observed by the gain of an electron forming an anion.

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Thus, the metallic character is highest at the extreme left and decreases while moving to the right of the periodic table; the non-metallic character increases while moving from left to right across the period. The chemical reactivity of an element can be best explained by its reactions with oxygen and halogens. Elements on two extremes of a period easily combine with oxygen to form oxides. The normal oxide formed by the element on extreme left is the most basic (e.g., Na<sub>2</sub>O), whereas that formed by the element on extreme right is the most acidic (e.g., Cl<sub>2</sub>O<sub>7</sub>). Oxides of elements in the centre are amphoteric (e.g., Al<sub>2</sub>O<sub>3</sub>, As<sub>2</sub>O<sub>3</sub>) or neutral (e.g., CO, NO, N<sub>2</sub>O). Amphoteric oxides behave as acids with bases and as bases with acids, whereas neutral oxides have no acidic or basic properties.

**Problem 6:** Show by a chemical reaction with water that Na<sub>2</sub>O is a basic oxide and Cl<sub>2</sub>O<sub>7</sub> is an acidic oxide.

**Solution:** Na<sub>2</sub>O with water forms a strong base whereas Cl<sub>2</sub>O<sub>7</sub> forms strong acid.



Their basic or acidic nature can be qualitatively tested with litmus paper.

Among transition metals (3d series), the change in atomic radii is much smaller as compared to those of representative elements (s & p block) across the period. The change in atomic radii is still smaller among inner -transition metals (4f series). The ionisation enthalpies are intermediate between those of s- and p-block elements. Therefore, transition metals are less electropositive than group 1 and 2 metals. In a group, the increase in atomic and ionic radii with increase in atomic number generally results in a gradual decrease in ionisation enthalpies and a regular decrease [with exception in some third period elements as shown in section 2 (d)] in electron gain enthalpies. Thus, the metallic character increases down the group and non-metallic character decreases. For transition elements, a reverse trend is observed because of the small change in atomic radii and ionisation enthalpy.

#### 4. Summary

The module deals with the periodic trends in atomic radii, ionisation enthalpies, electron gain enthalpies, electronegativity, valency metallic acidic and basic character. The atomic radii decrease while going from left to right in a period and increase with atomic number in a group. Ionisation enthalpy and electronegativity generally increase across a period and decrease down a group. Electron gain enthalpies, in general, become more negative across a period and less negative down a group. There is some periodicity in valency, for example, among representative elements; the valency is either equal to the number of electrons in the

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outermost orbitals or eight minus this number. Chemical reactivity is highest at the two extremes of a period and is lowest in the centre. The reactivity on the left extreme of a period is because of the ease of electron loss (or low ionisation enthalpy). Oxides formed of the elements on the left are basic and of the elements on the right are acidic in nature. Oxides of elements in the centre are amphoteric or neutral.