

CLASSIFICATION OF ELEMENTS & PERIODICITY IN PROPERTIES

MODERN PERIODIC TABLE (MODIFIED MENDELEEV'S PERIODIC TABLE):

- (i) It was proposed by Moseley.
- (ii) Modern periodic table is based on atomic number.

Moseley's equation $\sqrt{\nu} = a(Z - b)$, ν = frequency of x-rays, a & b = Moseley's constant, Z = Atomic number of element.

$$\sqrt{\nu} \propto Z$$

- (iii) **Modern periodic law** : The physical & chemical properties of elements are the periodic function of their atomic number.

(iv) Nomenclature of elements :

- (a) IUPAC gave names to elements above atomic number 100 as follows –

0	1	2	3	4	5	6	7	8	9
nil	un	bi	tri	quad	pent	hex	sept	oct	enn

- (b) In all the elements suffix is – ium. **Ex.**

Atomic No.	IUPAC Name	Symbol	Elemental Name	Symbol
101	Un nil unium	Unu	Mendelevium	Md
104	Un nil quadium	Unq	Rutherfordium	Rf
110	Un un nilium	Uun	Darmstadtium	Ds

Valency

Valency with respect to Hydrogen and Chlorine : Valency of H = 1 and Cl = 1.

It is defined as the number of hydrogen or Chlorine atoms attached with a particular element.

	IA	IIA	IIIA	IVA	VA	VIA	VIIA
	NaH	MgH ₂	AlH ₃	SiH ₄	PH ₃	H ₂ S	HCl
	NaCl	MgCl ₂	AlCl ₃	SiCl ₄	PCl ₃	SCl ₂	Cl ₂
Valency	1	2	3	4	3	2	1

Note : Valency w.r.t. H across the period increases upto 4 and then again decreases to 1.

Valency with respect to oxygen : Valency of 'O' = 2

It is defined as twice the number of oxygen atoms attached with a particular atom.

	IA	IIA	IIIA	IVA	VA	VIA	VIIA
	Na ₂ O	MgO	Al ₂ O ₃	SiO ₂	P ₂ O ₅	SO ₃	Cl ₂ O ₇
Valency	1	2	3	4	5	6	7

Note : Valency with respect to oxygen increases from 1 to 7 across the period. Valency w.r.t. 'O' is equal to the group number.

Atomic Radius

The average distance of valence shell e^- from nucleus is called atomic radius.

Based on the chemical bonds, atomic radius is divided into four categories –

- (A) Covalent radius (B) Ionic radius (C) Metallic radius (D) Vander waal radius

Ionic Radius

(i) Cationic Radius

- (a) Cationic radius is always smaller than atomic radius **because** after loosing e^- number of e^- reduces, but number of protons remains same, due to its Z_{eff} increases, hence electrons pulls towards nucleus and atomic radius decreases.

- (b) Size of cation $\propto \frac{1}{\text{Magnitude of the charge } e \text{ or } Z_{eff}}$

- Ex.** (i) $\text{Fe} > \text{Fe}^{+2} > \text{Fe}^{+3}$
 (ii) $\text{Pb}^{+2} > \text{Pb}^{+4}$
 (iii) $\text{Mn} > \text{Mn}^{+2} > \text{Mn}^{+3} > \text{Mn}^{+4} > \text{Mn}^{+5} > \text{Mn}^{+6} > \text{Mn}^{+7}$

(ii) Anionic Radius

Size of iso electronic species :

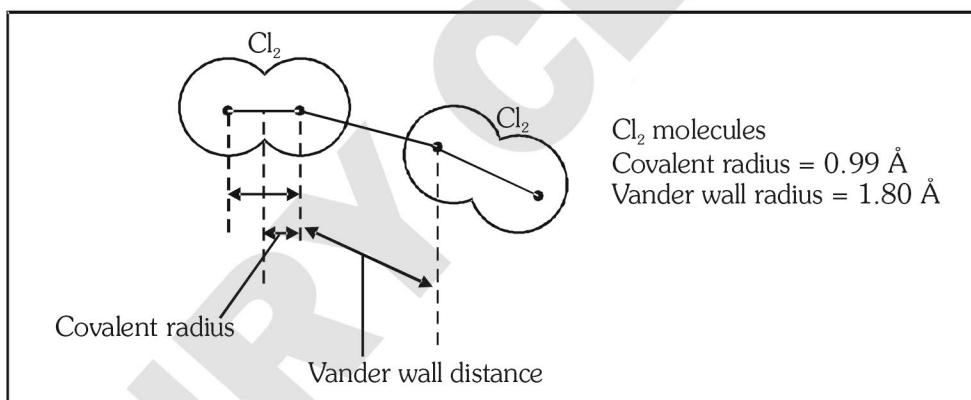
- Order of radius :** Ex.-($\text{S}^{2-} > \text{Cl}^- > \text{Ar} > \text{K}^+ > \text{Ca}^{+2}$), Ex.-($\text{N}^{3-} > \text{O}^{2-} > \text{F}^- > \text{Ne} > \text{Na}^+ > \text{Mg}^{+2} > \text{Al}^{+3}$),
 Ex.- $\text{P}^{3-} > \text{S}^{2-} > \text{Cl}^-$

Metallic Radius

- (a) Half of the nuclear distance between two adjacent metallic atoms in crystalline lattice structure.
 (b) there is no overlapping of atomic orbitals so, Metallic radius $>$ Covalent radius

Vander Waal's Radius

- (a) Those atoms (like noble gases) which are not bonded with each other, experiences a weak attractive force to come nearer.
 (b) Half of the distance between the nuclei of adjacently placed atoms in solid state of a noble gas is Vander Waal's radius.
 (c) Vander Waal radius $>$ Covalent radius.
 (d) Inert gas have only Vander Waal radius.



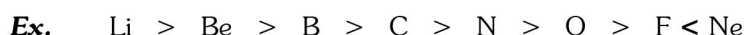
GOLDEN KEY POINTS

Factors affecting atomic size are :

- (a) Atomic radius $\propto \frac{1}{\text{Effective nuclear charge (Zeff)}}$
 $\text{Li} > \text{Be} > \text{B} > \text{C} > \text{N} > \text{O} > \text{F}$
- (b) Atomic radius \propto number of shells
 $\text{Li} < \text{Na} < \text{K} < \text{Rb} < \text{Cs}$
- (c) Atomic radius \propto Screening effect
- (d) Atomic size \propto Magnitude of $-ve$ charge
 $\text{O} < \text{O}^- < \text{O}^{2-}$
- (e) Atomic radius $\propto \frac{1}{\text{Magnitude of } +ve \text{ charge}}$
 $\text{Mn} > \text{Mn}^{+2} > \text{Mn}^{+3} > \text{Mn}^{+4}$
- (f) Atomic radius $\propto \frac{1}{\text{Bond order}}$
 $>\text{N} - \text{N}< > -\text{N} = \text{N} - > \text{N} \equiv \text{N}$

- **Periodic variation of atomic size :**

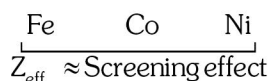
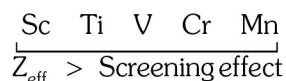
(i) **Across a period :** It decreases from left to right in a period as nuclear charge increases



(ii) **In a group :** It increases from top to bottom in a group as number of shell increases



Exceptions : Transition elements



Order of atomic radius –



- **Lanthanoid Contraction :**

In lanthanoid series Ce to Lu atomic size or ionic size regular decreases.

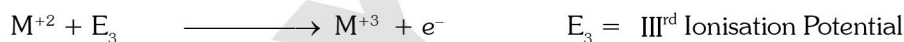
Ionisation Potential or Ionisation Energy or Ionisation Enthalpy

(i) Minimum energy required to remove most loosely bonded outer most shell e^- in ground state from an isolated gaseous atom is known as ionisation potential.

(Isolated \rightarrow Without any bonding with other atom)

(ii) **Successive Ionisation Energy**

(a) For an atom M, successive ionisation energies are as follows -



(b) Electron can not be removed from solid state of an atom, it has to convert in gaseous form, Energy required for conversion from solid state to gaseous state is called Sublimation energy.

(c) Ionisation Potential is always an endothermic process ($\Delta H = +ve$)

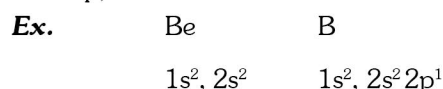
(d) It is measured in eV/atom (electron volt/atom) or Kcal/mole or KJ/mole

- **Penetration power of sub shells**

(a) Order of attraction of subshells towards nucleus (Penetration power) is -



(b) As subshell is more closer to nucleus so more energy will be required to remove e^- from s-subshell in comparison to p,d & f.



After losing one e^- , B attains electronic configuration of Be, so II^{nd} ionisation potential of B is more than Be.

(c) I.P. of full filled and half filled orbital is greater.

IInd Ionisation Potential of $B > Be$ **Stability of half filled and fully filled orbitals :**

Half filled p^3, d^5, f^7 or fully filled s^2, p^6, d^{10}, f^{14} are more stable than others so it requires more energy.

Ex.	N	O
	$1s^2, 2s^2 2p^3$	$1s^2, 2s^2 2p^4$

Ist Ionisation Potential order is $O < N$

Oxidation state :

Ionisation Potential \propto oxidation state of an atom

Ion with high oxidation state will have high ionisation potential. **Ex.** $Fe^{+3} > Fe^{+2} > Fe$

(iii) In a group : Size increase so ionisation potential decrease and along a period increases.

Li Na K Rb Cs

—————→

Size increases, Ionisation Potential decreases

Exception :

- Ionisation Potential $Ga > Al$ (While Ionisation Potential decreases down the group it is due to Transition contraction)

- Ionisation potential $Hf > Zr$ } (While ionisation potential should decrease down
5d 4d } the group. It is due to lanthanide contraction)

(iv) In a period : In a period atomic size decreases and z_{eff} increases so removal of electron becomes difficult and ionisation potential increases.

Li Be B C N O F Ne

—————→

atomic size decreases, z_{eff} increases, Ionisation Potential increases.

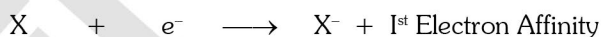
Order of ionisation potential $Li < B < Be < C < O < N < F < Ne$

Electron Affinity/Electron Gain Enthalpy

- (1) The amount of energy released or absorbed when electron is added to the valence shell of an isolated gaseous atom known as Electron affinity.



- (2) Mostly energy is released in the process of first Electron affinity.

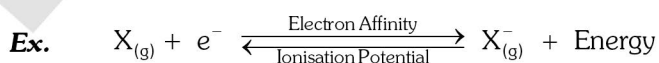


- (3) If Ist Electron affinity is exothermic

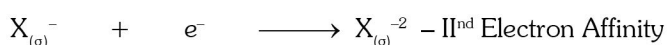
$$\Delta_{eg} H (\text{electron gain enthalpy}) = -ve \qquad \text{I}^{\text{st}} \text{Electron affinity is mostly positive}$$

- (4) Electron affinity of neutral atom is equal to ionisation potential of its anion.

{Energy evolved in Electron affinity to convert neutral atom to anion} = {Energy absorbed to remove e^- from anion}



- (5) On adding 2nd electron in anion X^-



or $X_{(g)}^- + e^- + \text{II}^{\text{nd}} \text{Electron Affinity} \longrightarrow X_{(g)}^{2-} \Delta_{eg} H (2^{\text{nd}} \text{electron gain enthalpy}) = +ve$

(6) **Factors affecting electron affinity :**

(A) Atomic size $\text{Electron Affinity} \propto \frac{1}{\text{Atomic size}}$

(B) Screening effect $\text{Electron Affinity} \propto \frac{1}{\text{Screening effect}}$

(C) Effective nuclear charge (Z_{eff}) $\text{Electron Affinity} \propto Z_{\text{eff}}$

 (D) **Stability of completely filled or half filled orbitals :**

Electron affinity of full filled or half filled orbital is very less or zero.

 (7) **In period :**

Electron affinity increases along the period due to increase in z_{eff} and decrease in atomic size.

Exception :

Cl > F > Br > I	S > O > P > N
S > Se > Te > O	Si > C > P > N

Electronegativity (EN)

(1) The tendency of an atom to attract shared electrons pair towards itself is called electronegativity.

Note : Small atoms are normally having more electronegativity than larger atoms.

 (2) **Factors Affecting electronegativity :**

 (A) **Atomic size**

$$\text{Electronegativity} \propto \frac{1}{\text{Atomic size}}$$

 (B) **Effective nuclear charge (Z_{eff})**

$$\text{Electronegativity} \propto Z_{\text{eff}}$$

 (C) **Hybridisation state of an atom**

$$\text{Electronegativity} \propto \% \text{ s character in hybridised atom}$$

 (D) **Oxidation state**

$$\text{Electronegativity} \propto \text{oxidation state}$$

$$sp > sp^2 > sp^3$$

$$Mn^{+2} < Mn^{+4} < Mn^{+7}$$

$$\text{s character } 50\% \quad 33\% \quad 25\%$$

$$O^{-2} < O^{-1} < O < O^{+1} < O^{+2}$$

$$\text{Electronegativity } 3.25 \quad 2.75 \quad 2.5$$

$$Fe < Fe^{+2} < Fe^{+3}$$

 (3) **Periodic table & Electronegativity :**

(A) Electronegativity decreases down the group.

(B) In period on moving from left to right electronegativity increases.

(C) In group of IIB elements (Zn, Cd, Hg) value of electronegativity increases down the group, because of lanthanide contraction

 (D) In IIIA group, value of electronegativity increases down the group, because of transition contraction
Electronegativity of Ga > Electronegativity of Al

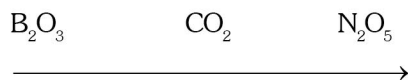
(a) Along a period acidic nature increases.

(b) Down the group basic nature increases

Note: BeO , Al_2O_3 , ZnO , SnO , PbO , SnO_2 , PbO_2 , Sb_2O_3 etc. are amphoteric oxides.

CO , H_2O , NO , N_2O etc. are neutral oxides.

Acidic strength of oxide and oxyacid \propto Electronegativity \propto O.S. or O.N.



EN increase, acidic nature increase.



(4) Electronegativity Scale :

(A) Pauling Scale (Bond energy scale) :

$$0.208\sqrt{\Delta_{AB}} = X_A - X_B \quad \text{Here } \Delta_{AB} \text{ is resonance energy of AB molecule}$$

This equation gives difference in electronegativity values.

(B) Mulliken scale : According to Mulliken electronegativity is average value of ionisation potential and electron affinity of an element,

$$X_m = \frac{\text{Ionisation Potential} + \text{Electron Affinity}}{2}$$

Relation between Pauling scale & Mulliken scale

$$X_p = 0.336 (X_m - 0.615)$$

X_p = Electronegativity given by Pauling

X_m = Electronegativity given by Mulliken

- If ionisation potential and electron affinity are given in eV, then electronegativity by Mulliken on Pauling scale will be

Pauling electronegativity is greater than Mulliken by 2.8

$$X_p = \frac{\text{Ionisation Potential} + \text{Electron Affinity}}{5.6} = \frac{\text{I.P.} + \text{E.A.}}{2 \times (2.8)} = \frac{X_m}{2.8}$$