

Chapter-1

Periodic Classification

Periodic Table

It may be defined as the table which classifies all the known elements in accordance of their properties in such a way that the elements having same properties are grouped together in the same vertical column.

Need for the Classification

With the discovery of large number of elements, it became difficult to study the elements individually, so classification was done to make the study easier.

Genesis of Periodic Classification

1 Lavoisier's Classification :-

- It was based on the basis of physical properties of the elements such as hardness, malleability and lustre.
- He classified the elements into two groups, i.e., metals and non metals.

Limitation:

- No place for metalloids.

2 Proust's Classification :-

- He assumed that all the elements were made up of hydrogen.

$$\text{Atomic weight of an element} = \frac{n \times (\text{mass of hydrogen})}{\text{number of H-atoms}}$$

Limitations:

- Fractional atomic masses.
- Elements are not made up of hydrogen.

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Dobereiner's Triads :-

→ Dobereiner classified the elements into groups of three elements with similar properties in such a manner so that the atomic weight of the middle element was the arithmetic mean of the other two elements.

Examples :

Element	At. no.	At. weight
Li	3	7
Na	11	23
K	19	39

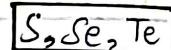
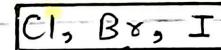
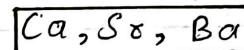
$\text{At. weight of Na} = \frac{\text{At. weight of Li} + \text{K}}{2} = \frac{7 + 39}{2} \Rightarrow 23$

Here, difference between atomic number of 'Li' and 'Na' = difference between atomic number of 'Na' and 'K'.

$$11 - 3 = 8$$

$$19 - 11 = 8$$

Other Triads :



Alkali earth metals

Halogens

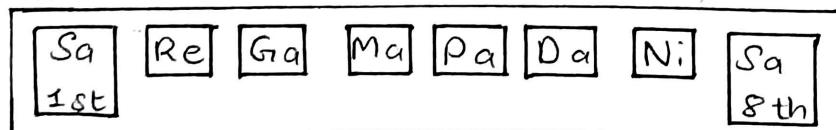
Limitation:

→ It was helpful in grouping some elements with similar characteristics together but he could not arrange all the elements known at that time into triads.

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Newland's Law of Octaves :-

- Newland (1865) found that if the elements were arranged in increasing order of their atomic weights, then properties of every eighth element were similar to those of first one.
- Newland called it law of octave.
- At that time 56 elements were known.
- Noble gases were not discovered at that time.
- It is like musical notes!



Example :

Sa , Re , Ga , Ma , Pa , Da , Ni						
Li	Be	B	C	N	O	F
Na	Mg	Al	Si	P	S	Cl
K	Ca					

Limitations

- In the above table, if we start counting from 'Li' then we get 'Na' as 8th element and we know that 'Li' and 'Na' shows similar properties.
- Similarly, if we start counting from 'B' then we get 'Al' as 8th element and we also know that 'B' and 'Al' shows similar properties.

Limitations:

- It was found that the law of octaves was applicable only upto calcium ($Z=20$) out of 56 known elements.
- Several new elements were discovered whose properties did not fit into the law of octaves.
- In order to fit elements into his table, Newland adjust two elements in the same slot. He also put some unlike elements under the same slot.

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Lothar Meyer's Curve

- He plotted a curve between atomic weight and atomic volume of different elements.
- Elements having similar properties occupy similar position in the curve.

The following observations can be made from the curve:

- Alkali metals (Li, Na, K, Rb, Cs, etc) occupy the peak positions on the curve.
- Alkaline earth metals (Be, Mg, Ca, Sr, Ba, etc) occupy the descending position on the curve.
- Most electronegative, i.e., halogens (F, Cl, Br, I) occupy the ascending position on the curve.
- Transition metals occupy bottom part of the curve.

Limitations:

- Meyer's periodic table was insufficient in comparison to Mendeleev's periodic table which was published the same year.
- There was no empirical or logical basis of classification and such values are difficult to remember.

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Mendeleev's Periodic Table

- The formulae of the hydrides and oxides formed by an element were treated as one of the basic properties of an element for its classification.

Mendeleev's Periodic Law :-

→ The physical and chemical properties of the elements are a periodic function of their atomic masses.

→ Elements with similar characteristics were present in vertical columns called groups.

→ The horizontal rows were known as periods.

→ This periodic table is divided into 7 horizontal rows (periods) and 8 vertical columns (groups) [I to VIII].

→ Zero group was added later on in the modified Mendeleev's Periodic Table.

→ Group I to VII (1-7) were divided into A and B.

Merits of Mendeleev's Periodic Table :-

→ Systematic study of the elements.

→ Prediction of new elements and their properties.

Eka-aluminium (Ga)

Eka-silicon (Ge)

Eka-boron (Sc)

Eka-manganese (Tc)

→ Mendeleev's periodic table helped in correcting the atomic masses of some of the elements like gold, beryllium and platinum based on their positions in the periodic table.

Demerits of Mendeleev's Periodic Table :-

- Position of hydrogen is not fixed because it shows properties similar to metals as well as non-metals.
- No separate position for isotopes.
- Some like elements were placed in different groups like Pt and Au.
- Some elements are not arranged in increasing order of their atomic masses.

Eg:-

Co placed before Ni,
As placed before K,
Te placed before I,
Th placed before Pa.

Periodic Law

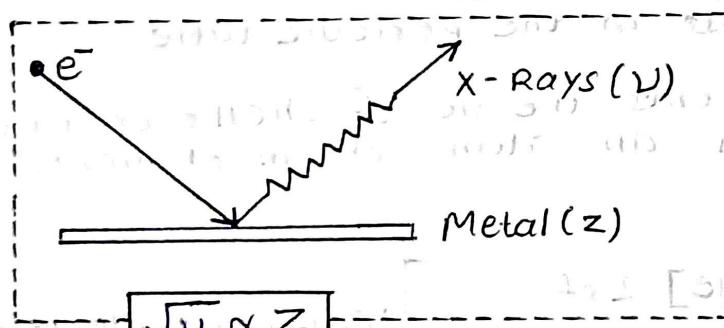
Anomalous Pairs

- Some unlike elements were placed in same group like Cu, Ag, Au were placed in 1st group with Na, K, etc. while they differ in their properties.

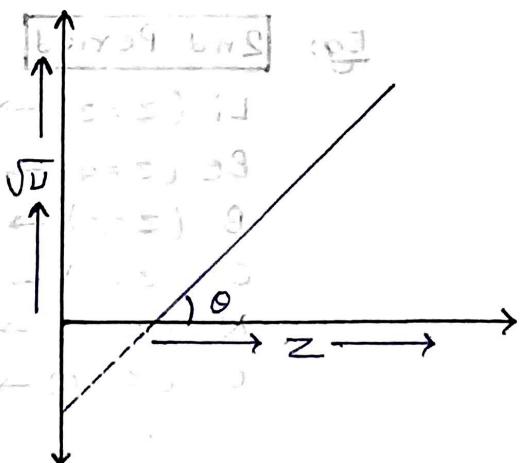
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Moseley's Periodic Law

- Moseley studied the frequency of X-ray produced by the bombardment of a strong beam of electrons on a metal target.
- He found that the square root of the frequency of X-ray is directly proportional to the atomic number of the metal.



$$\begin{aligned}\sqrt{\nu} &= az - ab \\ y &= mx + c\end{aligned}$$



Exact relation $\rightarrow \sqrt{\nu} = a(z-b)$

Also, $\sqrt{\nu} = az - ab$

z = Atomic no.
 b = constant
 a = constant

Modern Periodic Law :-

"Properties of elements are a periodic function of their atomic number."

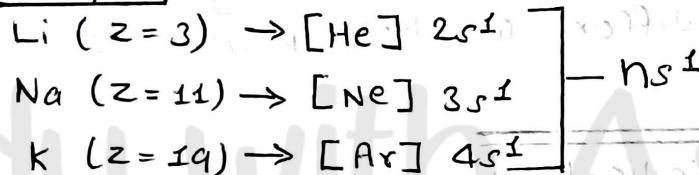
Periodic Function : Repetition of physical and chemical properties after a fixed interval. It is due to repetition of similar electronic configuration.

Modern Periodic Table

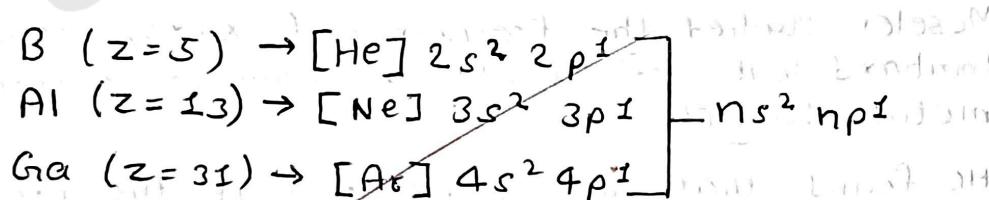
Groups in Modern Periodic Table :-

- There are 18 groups in the periodic table.
- Each group consists of elements having same outer shell electronic configuration.

Eg: **Group - I**



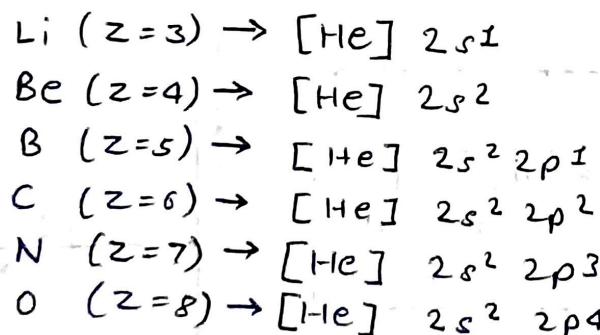
Group - 13



Periods in Modern Periodic Table :-

- There are 7 periods in the periodic table.
- Each period represents the no. of shells or energy levels present in an atom of an element.

Eg: **2nd Period**



"Outermost electrons of all elements in 2nd group are in 2nd shell"

Nomenclature of elements with atomic no > 100

Digit	Name	Abbreviation
0	nil	n
1	un	u
2	bi	b
3	tri	t
4	quad	q
5	pent	p
6	hex	h
7	sept	s
8	oct	o
9	enn	e

Examples:-

- 1) 101 \rightarrow Unnilunium (Unu)
- 2) 104 \rightarrow Unnilquadium (Unq)
- 3) 115 \rightarrow Ununpentium (Uup)
- 4) 118 \rightarrow Ununoctium (Uuo)
- 5) 114 \rightarrow Ununquadium (Uuq)
- 6) 119 \rightarrow Ununennium (Uue)
- 7) 109 \rightarrow Unnilennium (Une)

Note:

- Add 'ium' in last.
- Add only 'um' if name ends with 'i'.

Eg: - 103
(Unnil~~x~~trium)

Electronic Configuration of elements & Periodic Table

- \rightarrow S - subshell starts from 1st shell.
- \rightarrow P - subshell starts from 2nd shell.
- \rightarrow d - subshell starts from 3rd shell.
- \rightarrow f - subshell starts from 4th shell.

S - subshell = $2e^-$
P - subshell = $6e^-$
d - subshell = $10e^-$
f - subshell = $14e^-$

$$\begin{aligned} 1s &< 2s < 2p < 3s < 3p < 4s \\ 3d &< 4p < 5s < 4d < 5p < 6s \\ 4f &< 5d < 6p < 7s < 5f < 6d \\ 7p &< 8s \end{aligned}$$

Trick:

$$\begin{array}{l} (ss\ p)(sp\ s) \\ (dp\ s)\ (d\ ps) \\ (fd\ ps)(f\ dp\ s), \text{ old-2} \end{array}$$

Electronic Configuration in Periods:-

- 1st period \Rightarrow 2 elements $\rightarrow [1s] \rightarrow [z=1 \text{ to } z=2]$
- 2nd period \Rightarrow 8 elements $\rightarrow [2s, 2p] \rightarrow [z=3 \text{ to } z=10]$
- 3rd period \Rightarrow 8 elements $\rightarrow [3s, 3p] \rightarrow [z=11 \text{ to } z=18]$
- 4th period \Rightarrow 18 elements $\rightarrow [3d, 4s, 4p] \rightarrow [z=19 \text{ to } z=36]$
- 5th period \Rightarrow 18 elements $\rightarrow [4d, 5s, 5p] \rightarrow [z=37 \text{ to } z=54]$
- 6th period \Rightarrow 32 elements $\rightarrow [4f, 5d, 6s, 6p] \rightarrow [z=55 \text{ to } z=86]$
- 7th period \Rightarrow 32 elements $\rightarrow [5f, 6d, 7s, 7p] \rightarrow [z=87 \text{ to } z=118]$

Note: Highest value of 'n' gives the period no. of an element.

Eg: $1s^2 2s^2 2p^6$ [2nd Period]

$1s^2 2s^2 2p^6 3s^1$ [3rd Period]

$1s^2 2s^2 2p^6 3s^2 3p^1$ [3rd Period]

$2s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^1$ [4th Period]

1st Period \rightarrow Very short period (2 elements)

2nd Period \rightarrow Very short period (8 elements)

3rd Period \rightarrow short period (8 elements)

4th Period \rightarrow Long period (18 elements)

5th Period \rightarrow Long period (18 elements)

6th Period \rightarrow very long period (32 elements)

7th Period \rightarrow very long period (32 elements)

2	8	8	18	18	32	32
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Magic No. $2, 8, 18, 32$ \rightarrow half-filled shells

Electronic configuration in Groups :-

→ Elements of a group have the similar valence shell electronic configuration, same no. of valence electrons and similar properties.

Classification of elements in S, P, d and f block

S-block elements :-

- They are soft metals with low melting and boiling point.
- They are all metals.
- Reactivity of metals increases down the group.
- Good conductors of heat and electricity.
- They have low ionisation enthalpy.

Group-1: H ($z=1$)

Li ($z=3$)

Na ($z=11$)

K ($z=19$)

Rb ($z=37$)

Cs ($z=55$)

Fr ($z=87$)

General electronic config.

ns^1

Group-2: Be ($z=4$)

Mg ($z=12$)

Ca ($z=20$)

Sr ($z=38$)

Ba ($z=56$)

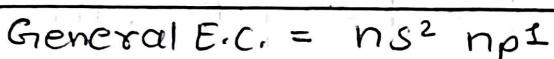
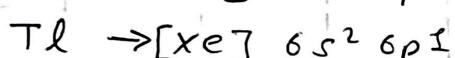
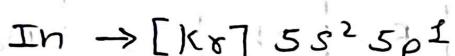
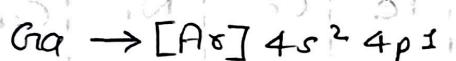
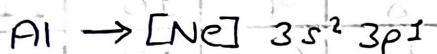
Rg ($z=88$)

ns^2

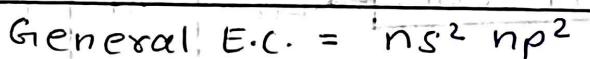
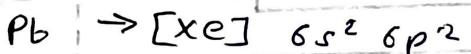
p-block elements :-

- It contains metals, nonmetals and metalloids.
- They mostly form covalent compounds.
- Ionisation energy is higher than s-block elements.
- These elements along with s-block elements are known as representative elements or main group elements.

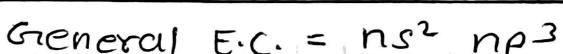
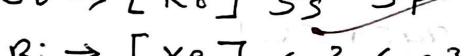
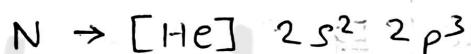
Group - 13 (Boron family)



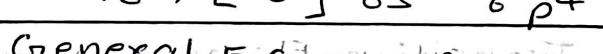
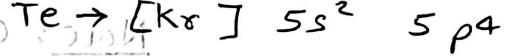
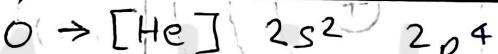
Group - 14 (carbon family)



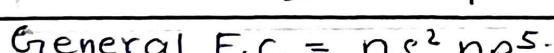
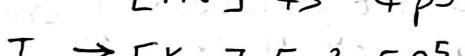
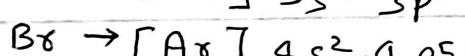
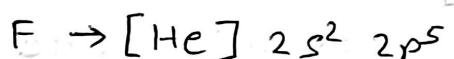
Group - 15 (Nitrogen family) (Pictogens)



Group - 16 (Oxygen family) (Chalogens - ore forming)



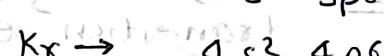
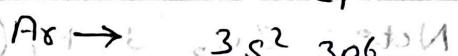
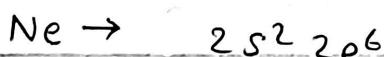
Group - 17 (Halogens) (salt forming)



General E.C.

$ns^{2-} np^{1-6}$

Group - 18 (Inert gases or Noble gases)

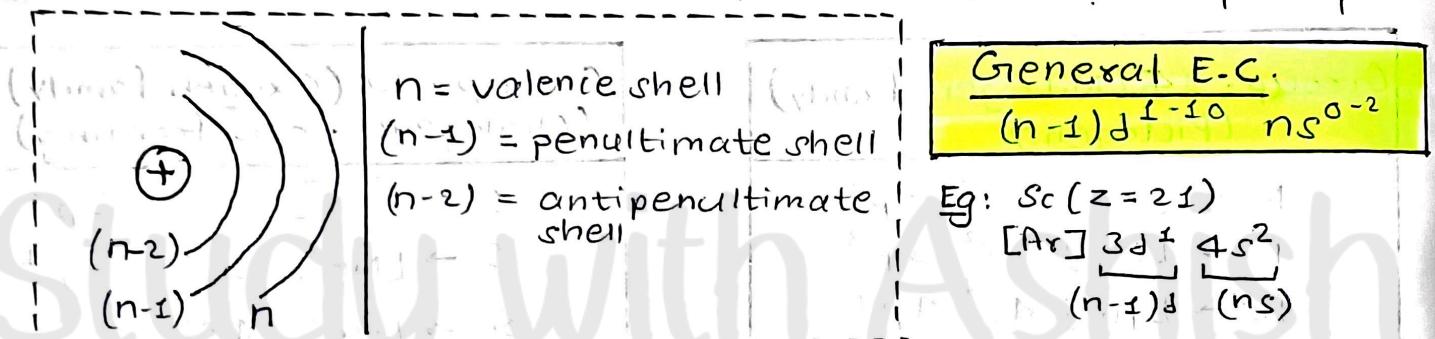


General E.C. = $ns^2 np^6$ (Ne-Rn)
= $1s^2$ (for He)

d-block elements :-

- These are hard, ductile and malleable metals with high melting and boiling point.
- They are good conductors of heat and electricity.
- They form coloured compounds and are used as catalysts.

Group →									
3	4	5	6	7	8	9	10	11	12
3d series → Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn
4d series → Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd
5d series → La	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg



Note: Cu, Ag, Au → Coinage metals

Transition Elements: Those elements which have incompletely filled d-orbitals in their ground state or any other oxidation state.



Here, +2 is the oxidation state of Zn in Zn^{+2} .

Note: Zn, Cd, Hg (Group - 12) are not considered as transition elements. Because of their d^{10} configuration. They are also called 'Pseudo Transition Elements'.

- All transition elements are d-block elements but all d-block elements are not transition elements (Zn, Cd, Hg)
- 3d-series ($z=21$ to $z=30$)
- 4d-series ($z=39$ to $z=48$)
- 5d-series ($z=57$, $z=72$ to $z=80$)

f-block Elements :-

- 1) Lanthanoids (4f series)
From $Z = 58$ to $Z = 71$ (14 elements)
 - 2) Actinoids (5f series)
From $Z = 90$ to $Z = 103$ (14 elements)
- They are heavy metals with high melting and boiling point.
- Most of the elements of Actinoid series are radioactive.
- Elements after Uranium are called as "transuranium elements".
- f-block elements are called inner-transition element

$$\text{General E.C.} = (n-2)f^{1-14} (n-1)d^{0-1} ns^2$$

Lanthanoids
(4f)

Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu
Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr

Actinoids
(5f)

~~3rd Group~~

Note

- The only liquid elements at STP are Bromine (Br) and Mercury (Hg).
- Caesium (Cs), Rubidium (Rb), Francium (Fr) and Gallium (Ga) become liquid at or just above room temperature.

Prediction of Group, Block & Period of an Element

- Write the electronic configuration.
- Highest value of 'n' gives the value of period no.
- **Block** is decided by the subshell which receives the last electron. (Highest value of $n+l$).

→ For Group No.

S-Block element = No. of valence e^-

P-Block element = 10 + No. of valence e^-

d- Block element = no. of e^- in $n s$ + no. of e^- in $(n-1)d$

f- Block element = Lanthanoids ($Z = 58$ to $Z = 71$)

Period = 6

Group = 3

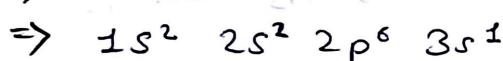
Actinoids ($Z = 90$ to $Z = 103$)

Period = 7

Group = 3

Examples :-

1) $Z = 11$

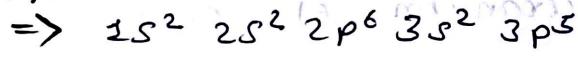


Period = 3

Block = s-Block

Group = 1

2) $Z = 17$



Period = 3

Block = p-block

Group = 17

3) $Z = 25$

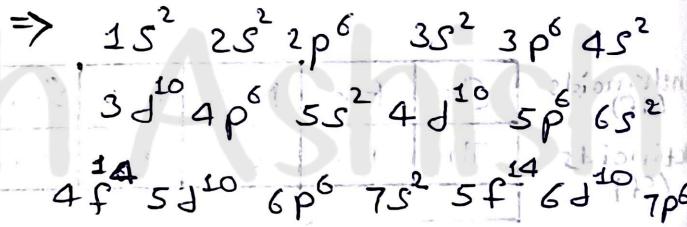


Period = 4

Block = d-block

Group = 7

4) $Z = 28$



⇒ Period = 7

Block = p-block

Group = 18

Classification of Metals, Non-metals & Metalloids

Metals :-

→ All s-block, d-block and f-block elements are metals.

→ Metals are usually solids at room temperature (Hg is an exception ; Ga and Ge also have very low melting points).

→ Metals usually have high melting and boiling point.

Metallic character increases down a group and decreases along a period as we move left to right.

Non-Metals :-

- Non-metals are usually solids or gases at room temperature with low melting and boiling points (B and C are exceptions).
- They are poor conductors of heat and electricity. Most non-metallic solids are brittle and are neither malleable nor ductile.

The non-metallic character increases as one goes from left to right across the periodic table and decreases down a group.

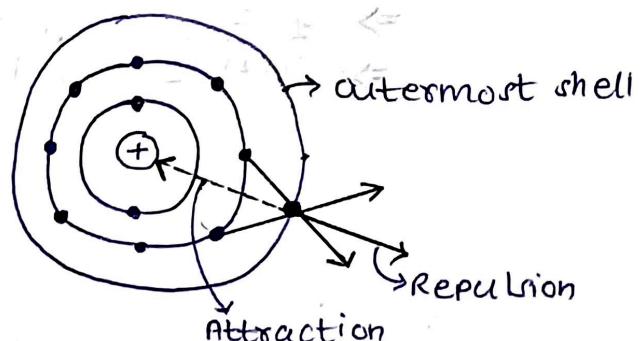
Metalloids :-

- Elements which lie along the borderline between metals and non-metals are called semimetals or metalloids.

Si, Ge, As, Sb, Te, Po and At

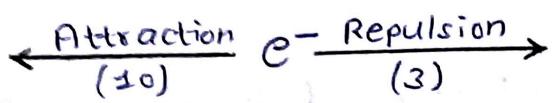
Shielding Effect

- Decrease in the net force of attraction on outermost e^- because of the presence of inner shell e^- is known as shielding effect or screening effect.
- It is the shielding of outermost e^- from nucleus by the inner shell e^- .
- Applicable for multi e^- species.
- Order of Shielding effect :- $S > P > d > f$
- d and f subshell electrons have poor shielding effect.
- As s and p are close to nucleus, so they are good shielders. Also s and p have simple structures.
- As d and f subshells have complex structure, the e^- in p and f are scattered. That's why d and f subshells are poor shielders.



Effective Nuclear Charge (Z_{eff})

→ Net force of attraction experienced by the outermost electron is known as effective nuclear charge (Z_{eff})



$$\text{Net Attraction} = \text{Attraction} - \text{Repulsion}$$
$$= 10 - 3 \Rightarrow 7 (Z_{\text{eff}})$$

Here, attraction is nuclear charge (z)

↓
Atomic No.
(No. of p)

$$\text{Effective Nuclear Charge (Z}_{\text{eff}}\text{)} = \frac{\text{Nuclear Charge (z)}}{\text{Repulsion (\sigma)}}$$

$$Z_{\text{eff}} = z - \sigma$$

σ = shielding/screening constant

Slater's Rule

It is used to calculate the value of shielding constant (σ) for an electron.

1) Write the electronic configuration of an atom as follows:

$$(1s) (2s, 2p) (3s, 3p) (3d) (4s, 4p) (4d) (4f) (5s, 5p)$$

Eg: Cl ($z = 17$)

$$\Rightarrow 1s^2 2s^2 2p^6 3s^2 3p^5$$

$$\Rightarrow (1s^2) (2s^2 2p^6) (3s^2 3p^5)$$

Mn ($z = 25$)

$$\Rightarrow 1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^5$$

$$\Rightarrow (1s^2) (2s^2 2p^6) (3s^2 3p^6) (3d^5) (4s^2)$$

2) For e⁻ present in s or p subshell

- (a) The e⁻ present on the right side of the group made by concerned e⁻, have no contribution.

Eg: $(1s^2)(2s^2 2p^6)(3s^2 3p^6) \underset{\substack{\downarrow \\ \text{concerned}}}{(4s^2)}$ \rightarrow no contribution.

- (b) The e⁻ present in the shell have the contribution of 0.35 each.

- (c) The e⁻ present in the $(n-1)$ shell have the contribution of 0.85 each.

- (d) The e⁻ present in the $(n-2)$ shell have the contribution of 1.0 each.

Eg: Ca ($Z = 20$) \rightarrow $(1s^2)(2s^2 2p^6) \underset{\substack{\downarrow \\ (\text{n-2}) \text{ shell}}}{(3s^2 3p^6)} \underset{\substack{\downarrow \\ (\text{n-1}) \text{ shell}}}{(4s^2)}$ same shell

$$\sigma = (1 \times 0.35) + (8 \times 0.85) + (10 \times 1)$$

$$\sigma = 0.35 + 10 + 8 = 17.15$$

$$Z_{\text{eff}} = Z - \sigma = 20 - 17.15 = 2.85$$

3) For e⁻ present in d or f subshell

- (a) The e⁻ present on the right side of the group made by concerned e⁻, have no contribution.

- (b) The e⁻ present in the same group have the contribution of 0.35 each.

- (c) The e⁻ present in left of this group have the contribution 1.0 each.

Eg: Fe ($Z = 26$) $\rightarrow 1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^5$
 $(1s^2)(2s^2 2p^6)(3s^2 3p^6) \underset{\substack{\downarrow \\ (\text{3d}^5)}}{(4s^2)} \underset{\substack{\downarrow \\ (\text{3d}^5)}}{(3d^5)} \underset{\substack{\downarrow \\ (\text{4s}^2)}}{(4s^2)}$

σ for $3d e^-$

$$\Rightarrow 5 \times 0.35 + 18 \times 1 = 19.75$$

$$Z_{\text{eff}} = 26 - (19.75) = 6.25$$

σ for $2s e^-$

$$\Rightarrow (7 \times 0.35) + (2 \times 0.85) = 4.2$$

$$Z_{\text{eff}} = 26 - 4.2 = 21.8$$

Note: $\begin{cases} \xrightarrow{\quad} \text{same group (in 3d rule)} \\ \xrightarrow{\quad} \text{single shell (for 2nd rule)} \end{cases}$

Periodicity in Properties

→ A regular increase or decrease in physical or chemical properties of elements with the change in atomic no., is known as periodicity.

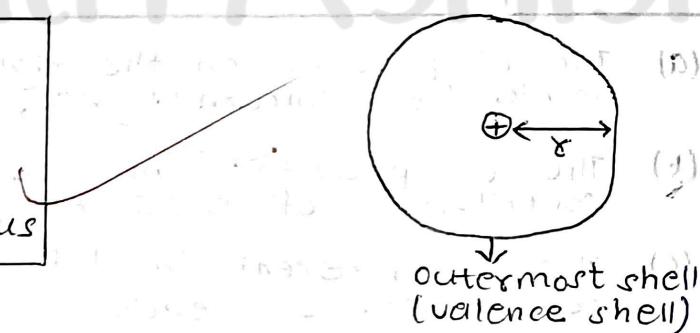
Periodicity in Physical Properties

- 1) Atomic Radius
- 2) Ionisation Enthalpy
- 3) Electron gain Enthalpy
- 4) Electronegativity

Atomic Radius

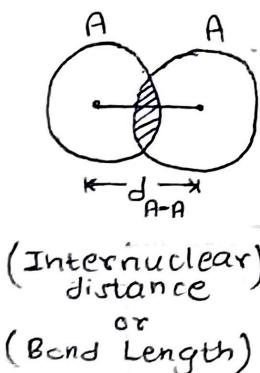
- Atomic Radius is the distance of outermost shell from the centre of nucleus for an isolated atom.
- Exact value of atomic radius cannot be calculated because it is not possible to get isolated atom (except for noble gases).
- Atomic radius is calculated in bonding state.

- Covalent Radius
Metallic Radius
Ionic Radius
Van-der waal Radius



1) Covalent Radius :-

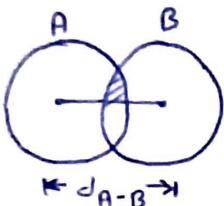
Case-I : For Homonuclear Molecules



Half of internuclear distance between two covalently bonded atoms is known as covalent radius.

$$r_{\text{cov.}} = \frac{d_{\text{A-A}}}{2}$$

Case - 2: For Heteronuclear Molecule



Internuclear distance

$$d_{A-B} = r_A + r_B - 0.09 (\Delta EN), \text{ Å}$$

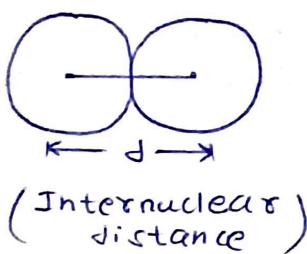
r_A : Covalent radius of A = $d_{AA}/2$

r_B : Covalent radius of B = $d_{BB}/2$

ΔEN : Difference in electronegativity

2) Metallic Radius :-

- It is one half the internuclear distance between the two metal ions in the metallic lattice.

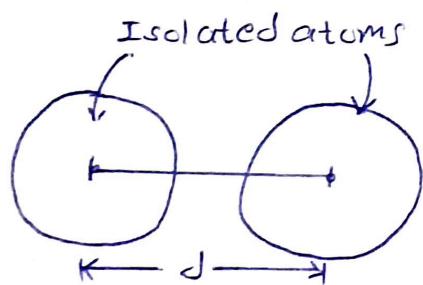


$$r_m = \frac{d}{2}$$

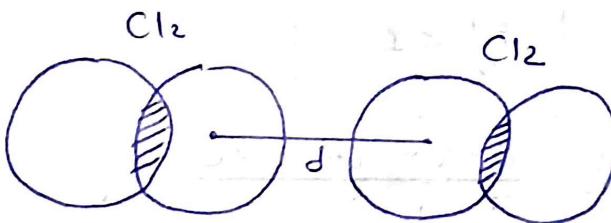
$$r_{\text{cov.}} < r_m$$

3) Van-der Waal's Radius :-

- It is one half the distance between the nuclei of two non-bonded isolated atoms or two adjacent atoms belonging to two neighbouring molecules of an element in the solid-state.



$$r_{\text{vander}} = \frac{d}{2}$$



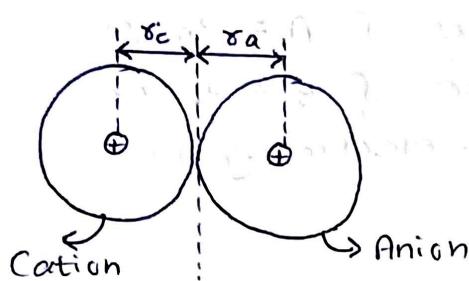
$$r_{\text{vander}} = \frac{d}{2}$$

- In the case of inert gases generally, van der waal's radius is defined so they have the largest size in their respective period.

$$r_{\text{covalent}} < r_{\text{metallic}} < r_{\text{vander}}$$

4) Ionic Radius :-

→ Ionic Bond → Cation + Anion
 (+ve) (-ve)



→ Size of cation is always smaller than its parent atom.

$$\begin{array}{|c|} \hline \text{Z/e ratio for Na:-} \\ \hline \frac{11}{10} = 1 \\ \hline \end{array}$$

$$\begin{array}{|c|} \hline \text{Na} > \text{Nat (size)} \\ \hline \downarrow & \downarrow \\ \text{Z=11} & \text{Z=11} \\ \text{P=11} & \text{P=11} \\ \text{e}^- = 10 & \text{e}^- = 10 \\ \hline \end{array}$$

$$\begin{array}{|c|} \hline \text{Z/e ratio for Nat} \\ \hline \frac{11}{10} > 1 \\ \hline \end{array}$$

Reason: With loss of e^- , Z_{eff} increases and size decreases.

→ Size of anion is always bigger than its parent atom.

$$\begin{array}{|c|} \hline \text{Z/e for Cl:-} \\ \hline \frac{17}{17} = 1 \\ \hline \text{Z/e for } \text{Cl}^-:- \\ \hline \frac{17}{18} < 1 \\ \hline \end{array}$$

$$\begin{array}{|c|} \hline \text{r}_{\text{Cl}} > \text{r}_{\text{Cl}^-} \text{ (size)} \\ \hline \text{Z=17} & \text{Z=17} \\ \text{P=17} & \text{P=17} \\ \text{e}^- = 17 & \text{e}^- = 18 \\ \hline \end{array}$$

Reason: With increase in e^- , Z_{eff} decreases and size of anion increases.

Eg: Order of size :-

$$(i) \text{A} > \text{A}^+ > \text{A}^{2+}$$

$$(ii) \text{B} < \text{B}^- < \text{B}^{2-}$$

$$(iii) \text{I}^+ < \text{I} < \text{I}^-$$

$$\frac{1}{2} = 0.5 \text{ m.u}^3$$

→ Isoelectronic species : Species having same no. of e^- :

Eg: Species = $\text{Na}^+ > \text{Mg}^{2+} > \text{Al}^{3+} > \text{F}^- < \text{O}^{2-} < \text{N}^{3-}$

$Z =$	11	12	13	9	8	7
$P_{Na} =$	11	12	13	9	8	7
$e^- =$	10	10	10	10	10	10
$Z/e^- =$	$\frac{11}{10} < \frac{12}{10} < \frac{13}{10}$			$\frac{9}{10} > \frac{8}{10} > \frac{7}{10}$		

- In case of isoelectronic species:
- Higher the +ve charge, smaller the size
 - Higher the -ve charge, larger the size

Factors affecting Atomic Radius

1) Number of Shells :-

$$\text{Atomic Radius} \propto \text{No. of shells.}$$

2) Effective Nuclear Charge (z_{eff}) :-

$$z_{\text{eff}} \propto \frac{1}{\text{size}(r)}$$

3) Screening / Shielding effect (σ) :-

$$\text{Atomic Radius} \propto \sigma$$

Periodic Variation of Atomic Radii

For Normal Elements (s and p block elements)]

→ In a period from left to right, z_{eff} increases, as a result size decreases.

Eg: 2nd Period Li > Be > B > C > N > O > F

3rd Period Na > Mg > Al > Si > P > S > Cl

Exception :

Li > Be > B > C > N > O > F > Ne Wrong

Ne > Li > Be > B > C > N > O > F Right

Note: Noble gases have the largest size in their respective period because van der waal radius is defined for them.

→ In a group no. of shell increases, z_{eff} remains constant, as a result atomic radius increases down the group.

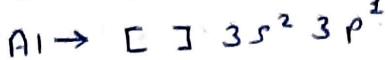
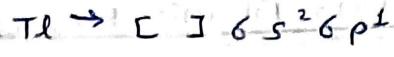
Eg: Li < Na < K < Rb < Cs
Be < Mg < Ca < Sr < Ba

Exception : Group-13 Elements :-



1)

$$\gamma_{Al} > \gamma_{Ga}$$

Energy $\rightarrow 3d < 4p$ $n=3 \quad n=4$ $l=2 \quad l=1$ $n+d=5 \quad n+l+5$  $n \uparrow \text{Energy} \uparrow$ Reason : Poor Shielding of 3d e-

Because of poor shielding of 4f e⁻, the force of attraction increases on outermost e⁻. As a result size decreases. This phenomenon is known as Lanthanide contraction.

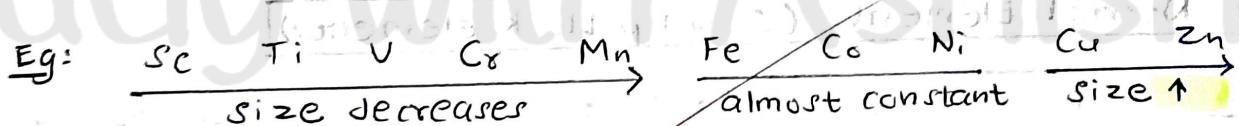
2)

$$\gamma_{In} \approx \gamma_{Tl}$$

Energy $\rightarrow 4f < 6p$ Reason : Poor shielding of 4f e- (Lanthanide contraction)

For Transition Elements (d-block Elements)

\rightarrow In a period size firstly decreases, after that it becomes constant and then increases.

Attraction $>$ Repulsion

Attraction = Repulsion

Attraction $<$ RepulsionReason : Poor shielding of 3d e⁻. \rightarrow In a group

Group - 3

Sc - 3d

Y - 4d

La - 5d

Ac - 6d

size increases down the group

From Group - 4 to 12

$$3d < 4d \approx 5d$$

Reason : Poor shielding of 4f e⁻ (Lanthanide contraction).Energy : $4f < 5d$

Inner Transition Elements (f block elements)

→ Ionic Radius: Size decreases from left to right.

Reason: 4f series (Lanthanoids)

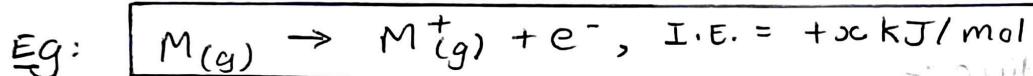
Poor shielding of $4f^{7-}$ (Lanthanide contraction)

5f series (Actinoids)

Poor shielding of $5f^{7-}$ (Actinoid contraction)

Ionisation Enthalpy / Ionisation Potential

→ It is the amount of energy required to remove e^- from valence shell of isolated gaseous atom in its ground state.



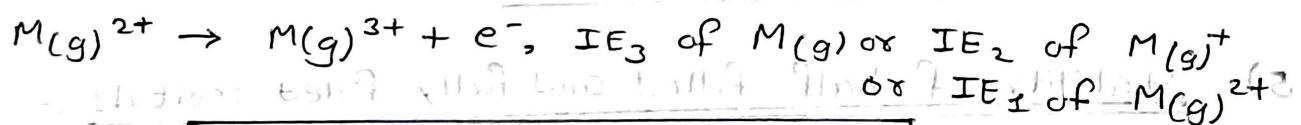
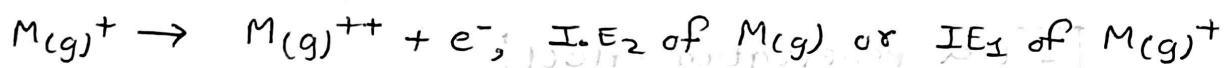
→ Ionisation energy is +ve that is it means it is always given from outside to remove e^- .

Endothermic process $\Delta H = +ve$

↓
Enthalpy change

→ Higher the ionisation energy, it is difficult to remove an e^- from the atom.

→ Successive Ionisation Energy:



I.E₁: 1st ionisation energy

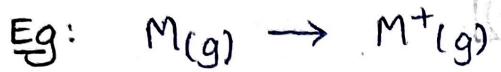
I.E₂: 2nd ionisation energy

I.E₃: 3rd ionisation energy

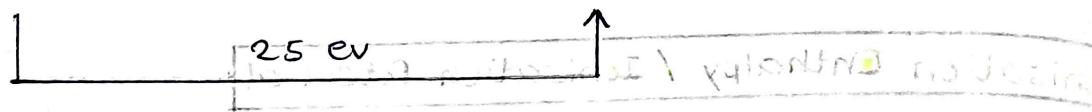
I.E₁, I.E₂, I.E₃, ... → Successive Ionisation Energy

$$I.E_1 < I.E_2 < I.E_3 < I.E_4 < \dots$$

Reason: It is difficult to remove from a +ve ion than a neutral atom.



Calculate - (i) IE₂ of $M(g)$ = 15 ev
(ii) IE₁ of $M(g)^+$ = 15 ev



Factors Affecting Ionisation Enthalpy

1) Atomic Radius :-

$$I.E \propto \frac{1}{\text{atomic radius}}$$

2) Effective Nuclear Charge :-

$$I.E \propto Z_{\text{eff}}$$

3) Shielding Effect (σ) :-

$$I.E \propto \frac{1}{\sigma}$$

4) Penetration effect of orbitals :-

Penetration $\rightarrow s > p > d > f$

$$I.E \propto \text{Penetration effect}$$

5) Stability of half filled and fully filled orbitals :-

Half filled and fully filled orbitals are more stable.

$$I.E \propto \text{stability of half filled and fully filled orbitals}$$

$$1s^1 > 2s^1 > 2p^3 > 3s^2$$

Periodicity in Ionisation Enthalpy

For Normal Elements (s and p block elements)

- In a period from left to right ionisation enthalpy generally increases because size decreases and z_{eff} increases.

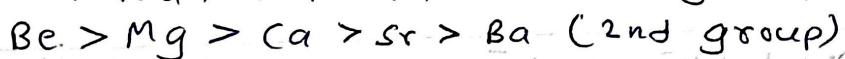
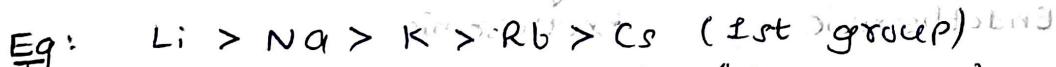
Exception: 1) 2nd Period elements:



Behaviour of Be & B	Behaviour of N & O	Ne
$\text{Be}(z=4) \rightarrow 1s^2 2s^2$ $\text{B}(z=5) \rightarrow 1s^2 2s^2 2p^1$ $I.E = \text{Be} > \text{B}$ <u>Reason:</u> Penetration of $2s e^-$ is more than that of $2p e^-$	$\text{N}(z=7) \rightarrow 1s^2 2s^2 2p^3$ Half filled $\begin{array}{ c c c }\hline & \uparrow & \uparrow \\ \hline & & 2p \\ \hline \end{array}$ (more stable) $\text{O}(z=8) \rightarrow 1s^2 2s^2 2p^4$ $\begin{array}{ c c c }\hline & \uparrow & \uparrow \\ \hline & & 2p \\ \hline \end{array}$ $I.E = \text{N} > \text{O}$ <u>Reason:</u> Stability of half filled orbital	$\text{Ne}(z=10)$ $1s^2 2s^2 2p^6$ $\begin{array}{ c c c }\hline & \uparrow & \uparrow \\ \hline & & 2p \\ \hline \end{array}$ (fully filled) (Highly stable) High I.E.

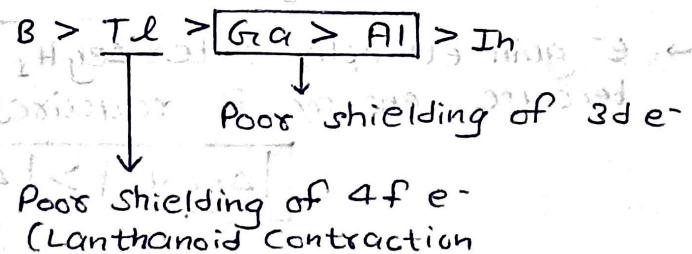
- 2) Same exception occurs in 3rd period due to the same reason as for 2nd period.
 $\text{Na} < \boxed{\text{Al}} < \boxed{\text{Mg}} < \boxed{\text{Si}} < \boxed{\text{S}} < \boxed{\text{P}} < \boxed{\text{Cl}} < \boxed{\text{Ar}}$

- In a group, atomic radius increases, as a result ionisation enthalpy generally decreases.



Exception: 1) Group - 13 Order of I.E. :-

B	$\delta_{\text{Al}} > \delta_{\text{Ga}}$
Al	
Ga	
In	$\delta_{\text{In}} \leq \delta_{\text{Ti}}$
Ti	



2) Group - 14 Elements :-

Order of I.E $\Rightarrow C > Si > Ge > Pb > Sn$

Poor shielding of 4f e-

For Transition Elements (d block elements)

\rightarrow In a group

For 3rd group

Sc

✓

Y

✓

La

✓

Ac

Sc > Y > La > Ac

For G-4 to G-12

$3d < 4d \approx 5d$ (size)

$3d > 4d$
 $5d > 4d$

(I.E)

\rightarrow In a period

General Trend : From Left to right

size \downarrow , I.E \uparrow

Electron Gain Enthalpy / Electron Affinity (ΔegH)

\rightarrow The amount of energy associated with the gain of e^- by an isolated gaseous atom in its ground state is known as e^- gain enthalpy.

Eg: $X(g) + e^- \rightarrow X(g)$, $\Delta egH_1 = (+ve/-ve)$

$\rightarrow \Delta egH$ can be +ve as well as -ve.
 ↓ Endothermic ↓ Exothermic

$X(g) + e^- \rightarrow X^-(g)$, $\Delta egH_1 = (+ve/-ve)$

$X^-(g) + e^- \rightarrow X^{2-}(g)$, $\Delta egH_2 = (+ve)$ Always

$\rightarrow e^-$ gain enthalpy after ΔegH_1 are always +ve (endothermic) because energy is required to overcome e^-e^- repulsion.

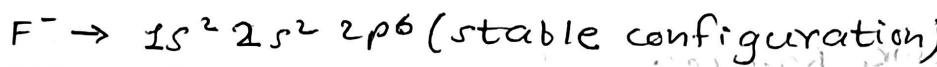
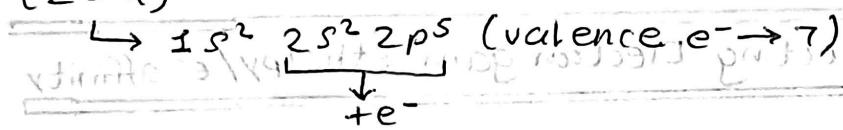
$|\Delta egH_2| > |\Delta egH_1|$

Negative Electron gain enthalpy ($\Delta egH = -ve$) (exothermic) :-

$$\text{Energy} \propto \frac{1}{\text{stability}}$$

Eg: Halogens (group-17 elements).

$$F(z=9)$$

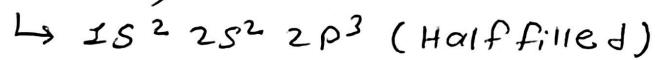


→ Jo e^- gain karna chahta hai $\rightarrow \Delta egH (-ve)$.

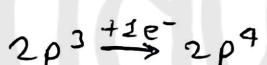
→ e^- affinity $\rightarrow +ve$

Positive Electron gain enthalpy ($\Delta egH = +ve$) (endothermic) :-

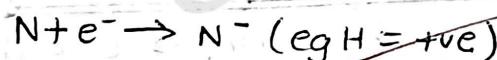
$$N(z=7)$$



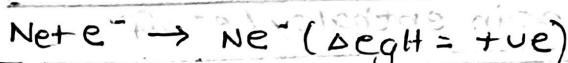
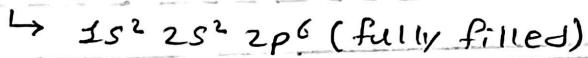
(more stable)



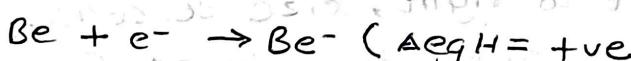
↓ (less stable)



$$Ne(z=10)$$



$$Be(1s^2 2s^2)$$



→ Be, N, Ne have +ve value of ΔegH because they do not have a tendency to gain e^- .

→ Jo e^- gain nahi karna chahta hai $\rightarrow \Delta egH (+ve)$

→ ΔegH gives an idea about the ease of e^- an element can gain.

→ More -ve $\Delta egH \Rightarrow$ utni aasani se e^- add hoga.

→ e^- affinity $\rightarrow -ve$

- Note:
- 1) Most of the elements have -ve $\Delta e.g.H.$
 - 2) These elements have +ve e^- -gain enthalpy -
 - Group-2 elements
 - Nitrogen (Half filled)
 - Noble gases (fully filled)

Factors Affecting Electron gain enthalpy/e- affinity

1) Atomic Radii :-

$\boxed{\text{Size} \uparrow, EA \downarrow}$

2) Effective nuclear charge :-

$\boxed{Z_{\text{eff}} \uparrow, EA \uparrow}$

3) Shielding effect (σ) :-

$\boxed{\sigma \uparrow, Z_{\text{eff}} \downarrow, EA \downarrow}$

4) Stability of half filled and fully filled orbitals :-

Half filled or fully filled \rightarrow more stable \rightarrow e^- gain nahi hoga.

$\boxed{EA \downarrow}$

Periodicity in Electron gain enthalpy/e- affinity

\rightarrow In a period, from left to right, size decreases, Z_{eff} increases, it becomes easier to gain e^- . Generally e^- affinity increases or e^- gain enthalpy becomes more negative.

Exception: 1) $\boxed{\text{Be, Mg, N, Noble gases}}$ \rightarrow +ve e^- gain enthalpy
 \rightarrow No e^- affinity or very less e^- affinity

$$2) [EA]_c > [EA]_N$$

$$[EA]_{Si} > [EA]_P$$

$$(G_i-14) \quad (G_i-15)$$

$$3) [EA]_{(G_i-1)} > [EA]_{(G_i-2)}$$

→ In a group, generally size increases, it becomes difficult to gain e^- , as a result e^- affinity decreases or e^- gain enthalpy becomes less -ve (with exceptions).

Exceptions: e^- affinity of 3rd period elements is higher than 2nd period elements.

ΔegH of 3rd period elements is more -ve than 2nd period elements.

Eg: O (2nd period)

S (3rd period)

$$EA \rightarrow S > O$$

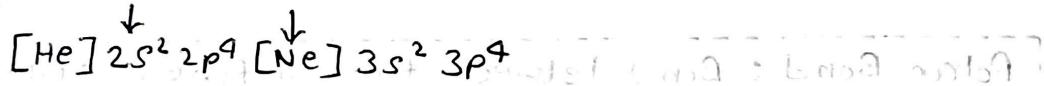
F (2nd period)

Cl (3rd period)

$$EA \rightarrow Cl > F$$

G-2 :	B	C	N	O	F
	Λ	Λ	Λ	Λ	Λ
G-3 :	Al	Si	P	S	Cl

Reason: O < S



Because of smaller size, elements of 2nd period faces more e^-e^- repulsion which is less in 3rd period elements because of their larger size.

→ Order of EA/ ΔegH for G-16 elements:-

O

S

Se

Te

Po

$$S > Se > Te > O$$

→ Block A elements

Block Z elements

→ Order of EA/ ΔegH for G-17 elements:-

F

Cl

Br

$$Cl > F > Br > I$$

→ EA of Cl is greatest in periodic table.

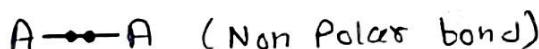
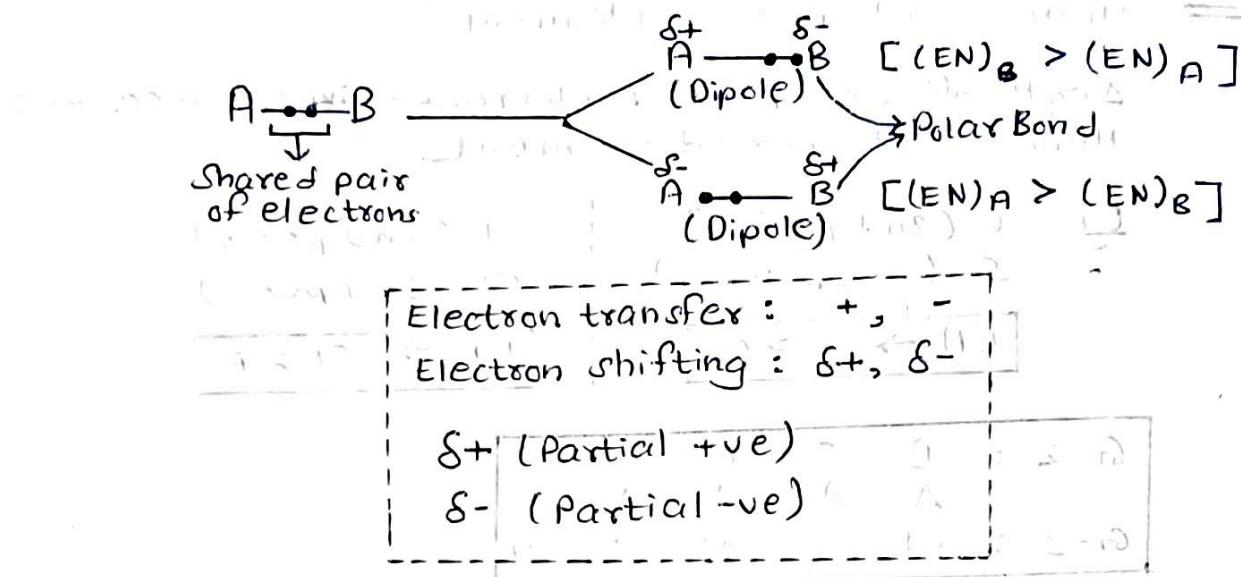
$$\text{Size } S-A \text{ To express bond } = S-A^3$$

$$\text{Size } H-A \text{ To express bond } = H-A^3$$

$$\text{Size } O-A \text{ To express bond } = O-A^3$$

Electronegativity

- It is defined as the tendency of an atom to attract the shared pair of electrons towards itself.



Polar Bond : Bond between two different atoms.

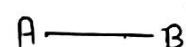
Non-Polar Bond : Bond between two same atoms.

- Electronegativity is calculated in Bonding state.

- The electronegativity of any given element is not constant. It varies depending on the element to which it is bound.

Electronegativity Scale :-

a) Pauling's scale



$$(EN)_A \rightarrow X_A$$

$$(EN)_B \rightarrow X_B$$

$$\Delta = X_A - X_B = 0.102 \sqrt{E_{A-B} - \sqrt{E_{A-A} \times E_{B-B}}} \quad (\text{KJ mol}^{-1})$$

$$\Delta = X_A - X_B = 0.208 \sqrt{E_{A-B} - \sqrt{E_{A-A} \times E_{B-B}}} \quad (\text{Kcal mol}^{-1})$$

E_{A-B} = Bond energy of $A-B$ bond

E_{A-A} = Bond energy of $A-A$ bond

E_{B-B} = Bond energy of $B-B$ bond

b) Mulliken's Scale

$$x_M = \frac{IP + EA}{2}$$

(both expressed in electron volts)

$x_M = EN$ on Mulliken scale

IP = Ionisation Potential

EA = Electron Affinity.

→ Relation between Pauling and Mulliken scale :

$$x_p = 1.35 (x_M) \pm 1/2 - 1.37$$

→ Mulliken's values were about 2.8 times larger than the Pauling's values.

Factors Affecting Electronegativity

1) Atomic Radius :-

No. of shells ↓, Radius ↓, EN ↑

$$\text{Atomic Radius} \propto \frac{1}{EN}$$

2) Effective nuclear charge :-

$z_{\text{eff}} \uparrow, EN \uparrow$

$z_{\text{eff}} \downarrow, EN \downarrow$

$$EN \propto z_{\text{eff}}$$

3) Charge on the atom :-

$$EN \propto (+ve \text{ charge})$$

4) Hybridisation state of an atom :-

$$EN \propto \% \text{ s-character}$$

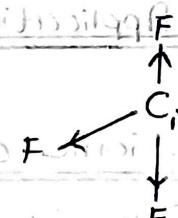
$$SP > SP^2 > SP^3 \rightarrow \text{Order of EN}$$

5) Nature of the bonded atom :-

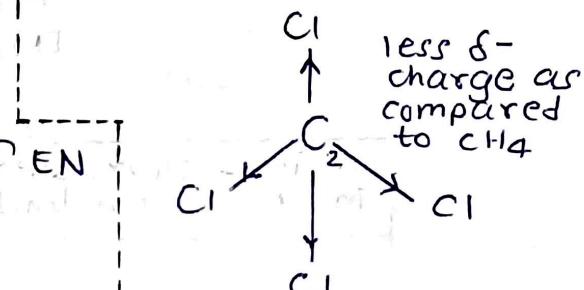
$$(EN)_C = 2.5 \quad (EN)_{Cl} > (EN)_C$$

$$(EN)_{Cl} = 3.0 \quad (EN)_F > (EN)_C$$

$$(EN)_F = 4.0 \quad (EN)_F > (EN)_{Cl}$$



more δ+ charge



less δ- charge as compared to CH4

$$\text{So, } (EN)_{C_1} > (EN)_{C_2}$$

Periodicity in Electronegativity

→ In a period from left to right, size decreases, Z_{eff} increases and electronegativity increases.

Most electronegative elements \rightarrow F
 Highly electronegative elements \rightarrow F, O, N \approx Cl

↓	↓	↓	↓
4	3.5	3	3

Eg: Li < Be < B < C < N < O < F

→ In a group from top to bottom, generally size increases and electronegativity decreases.

Exception: 1) In group -13 :-

B > Al < Ga < In < Tl

Al < Ga < In < Tl < B

2) In group -14 :-

C > Si > Ge \approx Sn \approx Pb

Important values of electronegativity :-

F > O > N \approx Cl > C \approx S \geq H

Applications of Electronegativity

1) Percentage ionic character in a covalent bond :-

→ None of the bond is 100% covalent or 100% ionic.



$$(\text{EN})_A \rightarrow x_A$$

$$(\text{EN})_B \rightarrow x_B \quad (\text{EN})_A > (\text{EN})_B$$

$$\% \text{ ionic character} = 16(x_A - x_B) + 3.5(x_A - x_B)^2$$

$$X_A - X_B = \Delta EN \text{ (difference in EN)}$$

$\Delta EN > 2.1$ [Ionic bond] \rightarrow Ionic character $> 50\%$

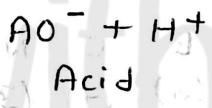
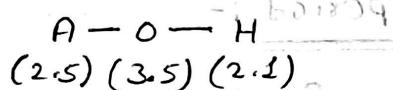
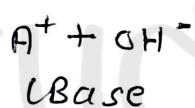
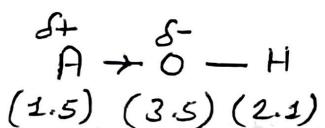
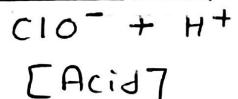
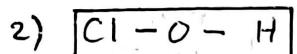
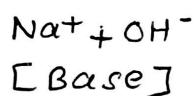
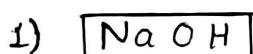
$\Delta EN < 2.1$ [Covalent bond] \rightarrow Ionic character $< 50\%$

2) Polar and Non polar bond :-

Polar Bond \rightarrow atoms of different electronegativity [A-B]

Non polar Bond \rightarrow atoms of same electronegativity [A-A]

3) Nature of Hydroxides :-



Trick

$\text{A}-\text{O}-\text{H}$ (acid)
If $(\text{EN})_{\text{A}} > 2.1$ [Acidic]
If $(\text{EN})_{\text{A}} < 2.1$ [Basic]

Valency or Oxidation state

\rightarrow The valency of representative elements is usually (not necessarily) equal to the no. of e^- in the outermost orbitals or equal to eight minus the no. of outermost e^- .

Case-I : No. of valence $e^- < 4$

Valency = No. of valence e^-
 $\text{Li} \rightarrow 1s^2 2s^1 = \text{valence } e^- = 1$
 valency = 1

$\text{B} \rightarrow 1s^2 2s^2 2p^1 = \text{valence } e^- = 3$
 valency = 3

Case-II: No. of valence $e^- > 4$

Valency = 8 - (No. of valence e^-)

$\text{N} \rightarrow 1s^2 2s^2 2p^3$

valence $e^- = 5$

valency = $8-5 = 3$

$\text{F} \rightarrow 1s^2 2s^2 2p^5$

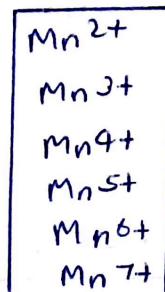
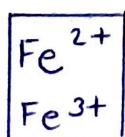
valence $e^- = 7$

valency = $8-7 = 1$

Group No.	1	2	13	14	15	16	17	18
Valence e ⁻	1	2	3	4	5	6	7	8
Valency	1	2	3	4	3, 5	2, 6	1, 7	0, 8

Note: d block elements can show variable valency because of a small energy gap between $(n-1)d$ and ns orbital

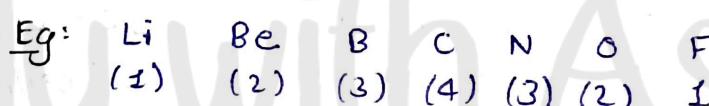
Eg:



Due to gap in energy $(n-1)d$ can also participate in bond formation.

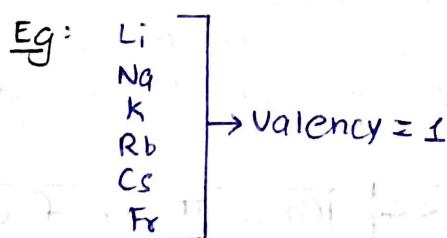
Periodicity along a period :-

→ Generally, valency first increases and then decreases.



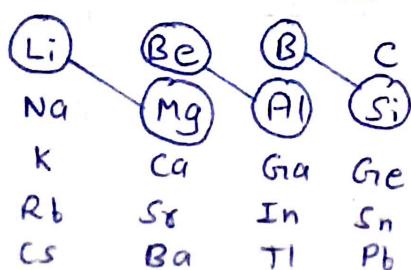
Periodicity in a group :-

→ Elements in a group have same no. of valence e⁻ and same valency.



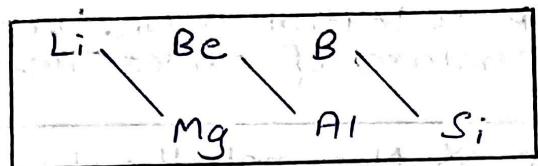
Anomalous properties of second period elements

→ The 1st element of each of the groups 1(Li) and 2(Be) and groups 13-17(B and F) differs in many respects from the other members of their respective group.



N O F (2nd period)

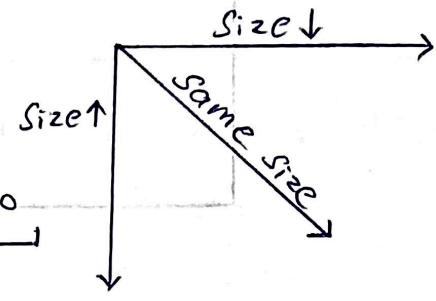
→ Diagonal Relationship: Similarity in properties of diagonally aligned elements.



Reason :-

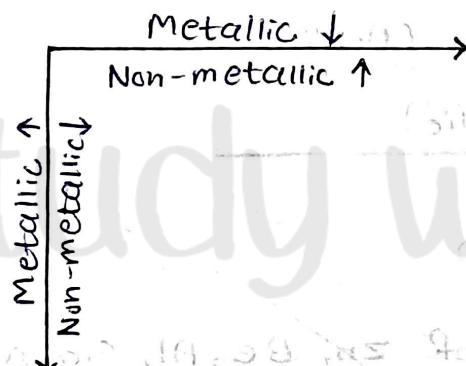
- 1) Similar size
- 2) Similar electronegativity
- 3) Similar (charge/radius) ratio

Polarising Power



Chemical Reactivity

1) Metallic and Non metallic Properties :-



Reactivity of metal : Tendency to lose e⁻

Reactivity of non metals : Tendency to gain e⁻

- Reactivity of metals (metallic character) increases down the group because it becomes easier to lose e⁻.
- Reactivity of non-metals (Non-metallic character) decreases down the group and increases in a period from left to right because it becomes difficult to gain e⁻ down the group.

Eg: Li < Na < K < Rb < Cs (Reactivity, metallic character)

F < Cl < Br < I (Reactivity, non-metallic character)

2) Nature of oxides :-

Metallic oxides → Basic
Non-metallic oxides → Acidic

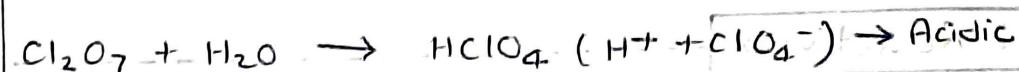
Na ₂ O	MgO	Al ₂ O ₃	SiO ₂	P ₂ O ₅	SO ₃	Cl ₂ O ₇
Highly Basic	Basic	Amphoteric	Slightly acidic	Acidic	Highly acidic	Very Highly Acidic

(3rd Period)

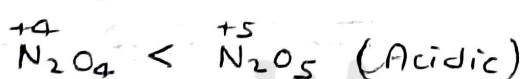
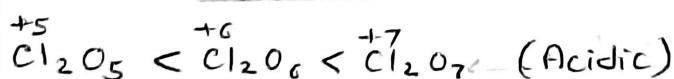
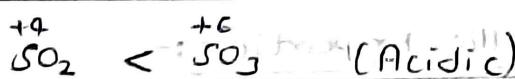
- In a period from left to right acidic nature of oxides increases and basic nature of oxides decreases.
- In a group from top to bottom acidic nature of oxides decreases and basic nature of oxides increases.

Acidic Nature & Non-metallic character
 & Electronegativity
 & +ve charge of the element
 (oxidation No.)

Proof:



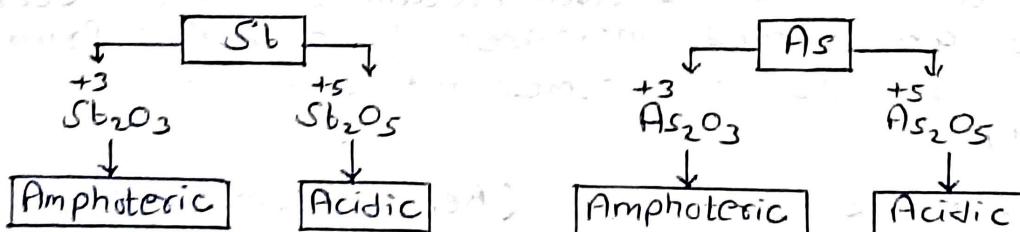
Eg:



→ Neutral Oxides: CO, NO, N₂O, H₂O

→ Amphoteric Oxides: All oxides of Zn, Be, Al, Ga, In, Pb are amphoteric oxides.

Trick: जलावै आती हो संत प्रवृत्त



Metals → +1, +2, +3 [Basic] Exception: Cr₂O₃ (Amphoteric)
 → +4 [Amphoteric]

→ +5, +6, +7 [Acidic] Exception: V₂O₅ (Amphoteric)

Non-Metals All oxides other than CO, NO, N₂O are acidic.

Metalloids Oxides of metalloids can be acidic or amphoteric.