PERIODIC TABLE

INTRODUCTION :

The arrangement of all the known elements according to their properties in such a way that the elements with similar properties are grouped together in a tabular form is called periodic table.

DEVELOPMENT OF PERIODIC TABLE

(A) LAVOISIER CLASSIFICATION :

(i) Lavoisier classified the elements simply in metals and non metals.

Metals are the one which have the tendency of losing the electrons.

$Na \rightarrow Na^+ + e^-$ and $K \rightarrow K^+ + e^-$

Non-metals are the one which have the tendency of gaining the electrons.

$F + e^- \rightarrow F^-$ and $Cl + e^- \rightarrow Cl^-$

(ii) **Drawback or Limitation :**

(a) As the number of elements increases, this classification became insufficient for the study of elements.

(b) There are few elements which have the properties of both metals as well as non-metals and they are called metalloids. Lavoisier could not decide where to place the metalloids.

(B) PROUT'S HYPOTHESIS :

He simply assumed that all the elements are made up of hydrogen, so we can say that

Atomic weight of element = n × (Atomic weight of one hydrogen atom)

$$\boxed{\text{Atomic weight of H} = 1}$$

where n = number of hydrogen atom = 1, 2, 3,....

Drawback or Limitation :

(i) Every element can not be formed by Hydrogen.

(ii) Atomic weight of all elements were not found as the whole numbers.

Ex. Chlorine (atomic weight 35.5) and Strontium (atomic weight 87.6)

(C) DOBEREINER TRIAD RULE [1817] :

(i) He made groups of three elements having similar chemical properties called TRIAD.

(ii) In Dobereiner triad, atomic weight of middle element is nearly equal to the average atomic weight of first and third element.

Ex.

Cl	Br	I	Ca	Sr	Ba	Li	Na	K
35.5	80.0	127	40	87.6	137	7	23	39

$$\left[x = \frac{35.5+127}{2} = 81.2 \right. \qquad \left[x = \frac{40+137}{2} = 88.5 \right. \qquad \left[x = \frac{7+39}{2} = 23 \right.$$

Where x=average atomic weight

(iii) Other examples – (K, Rb, Cs), (P, As, Sb), (S, Se, Te)

Drawback or Limitation : All the known elements could not be arranged as triads. It is not applicable for d and f-block elements.

(D) NEWLAND OCTAVE RULE [1865]

(i) He arranged the elements in the increasing order of their atomic mass and observed that the properties of every 8th element was similar to the 1st element. (like in the case of musical vowels notation)

(ii) At that time inert gases were not known.

Sa	Re	Ga	Ma	Pa	Dha	Ni	Sa
1	2	3	4	5	6	7	8

						H
Li	Be	B	C	N	O	F
Na	Mg	Al	Si	P	S	Cl
K	Ca					
↓	↓	↓	↓	↓	↓	↓

(iii) The properties of Li are similar to 8th element i.e. Na and Be are similar to Mg and so on.

Drawback or Limitation :

(a) This rule is valid only upto Ca because after Ca due to presence of d-block element there is a difference of 18 elements instead of 8 elements.

(b) After the discovery of Inert gas and including them in the periodic table, it has become the 8th element from Alkali metal so this law had to be dropped out.

(E) LOTHAR MEYER'S CURVE [1869] :

(i) He plotted a curve between atomic weight and atomic volume of different elements.

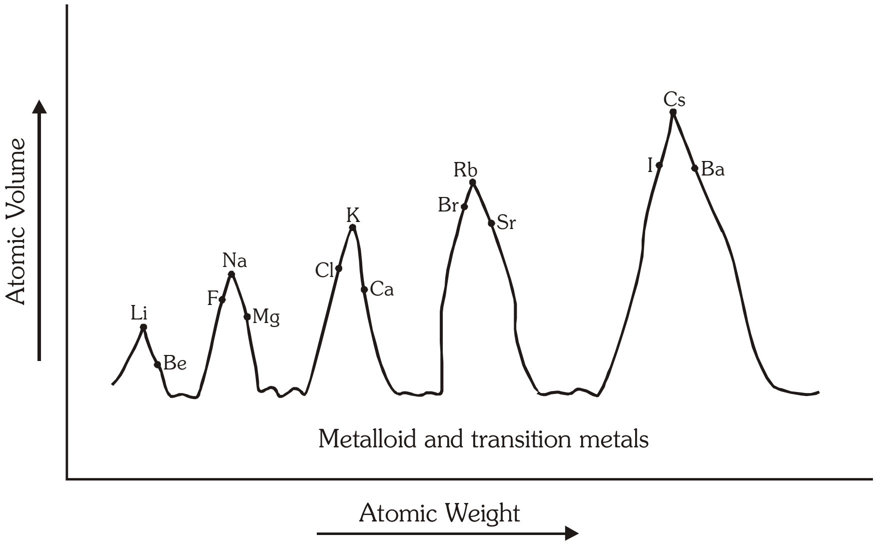
(ii) The following observation can be made from the curve –

- (a) Most electropositive elements i.e. alkali metals (Li, Na, K, Rb, Cs) occupy the peak positions on the curve.
- (b) Less electropositive i.e. alkaline earth metal (Be, Mg, Ca, Sr, Ba) occupy the descending position on the curve.
- (c) Metalloids (Si, Ge, As, Sb, Te, Po, At) and transition metals occupy bottom part of the curve.
- (d) Most electronegative i.e. halogens (F, Cl, Br, I) occupy the ascending position on the curve.

Note : Elements having similar properties occupy similar position on the curve.

Conclusion : On the basis of this curve Lother Meyer proposed that the physical properties of the elements are periodic function of their atomic weight and this has become the base of Mendeleev's periodic table.

Periodic function : When the elements are arranged in the increasing order of their atomic weight, elements having similar properties gets repeated after a regular interval.



(F) MENDELEEV'S PERIODIC TABLE [1869] :

(i) Mendeleev's periodic law : The physical and chemical properties of elements are the periodic function of their atomic weight.

(ii) Characteristics of Mendeleev's periodic table :

(a) It is based on atomic weight

(b) 63 elements were known, noble gases were not discovered.

(c) He was the first scientist to classify the elements in a systematic manner i.e. in horizontal rows and in vertical columns.

(d) Horizontal rows are called periods and there were 7 periods in Mendeleev's Periodic table.

(e) Vertical columns are called groups and there were 8 group in Mendeleev's Periodic table.

(f) Each group upto VII is divided into A & B subgroups.'A' sub group element are called normal or representative elements and 'B' sub group elements are called transition elements.

(g) The VIII group consisted of 9 elements in three rows (Transitional metals group).

(h) The elements belonging to same group exhibit similar properties.

(iii) Merits or advantages of Mendeleev's periodic table :

(a) Study of elements : First time all known elements were classified in groups according to their similar properties. So study of the properties of elements become easier .

(b) Prediction of new elements : It gave encouragement to the discovery of new elements as some gaps were left in it.

Sc (Scandium) Ga (Gallium) Ge (Germanium) Tc (Technetium)

These were the elements for whom position and properties were well defined by Mendeleev even before their discoveries and he left the blank spaces for them in his table.

Ex. Blank space at atomic weight 72 in silicon group was called Eka silicon (means properties like silicon) and element discovered later was named Germanium .

Similarly other elements discovered after mendeleev's periodic table were.

Eka Aluminium	– Galium(Ga)	Eka Boron	– Scandium (Sc)
Eka Silicon	– Germanium (Ge)	Eka Mangenese	– Technetium (Tc)

(c) Correction of doubtful atomic weights : Correction were done in atomic weight of some elements.

$$\boxed{\text{Atomic weight} = \text{Valency} \times \text{Equivalent weight.}}$$

Initially, it was found that equivalent weight of Be is 4.5 and it is trivalent (V = 3), so the weight of Be was 13.5 and there is no space in Mendeleev's table for this element. So, after correction, it was found that Be is actually bivalent (V = 2). So, the weight of Be became $2 \times 4.5 = 9$ and there was a space between Li and B for this element in Mendeleev's table.

Corrections were done in atomic weight of elements are – **U, Be, In, Au, Pt.**

(iv) Demerits of Mendeleev's periodic table :

(a) Position of hydrogen : Hydrogen resembles both, the alkali metals (IA) and the halogens (VIIA) in properties so Mendeleev could not decide where to place it.

(b) Position of isotopes : As atomic wt. of isotopes differs, they should have placed in different position in Mendeleev's periodic table. But there were no such places for isotopes in Mendeleev's table.

(c) Anomalous pairs of elements : There were some pair of elements which did not follow the increasing order of atomic weights.

Ex. Ar and Co were placed before K and Ni respectively in the periodic table, but having higher atomic weights.

$$\begin{pmatrix} \text{Ar} & \text{K} \\ 39.9 & 39.1 \end{pmatrix} \qquad \begin{pmatrix} \text{Te} & \text{I} \\ 127.5 & 127 \end{pmatrix}$$

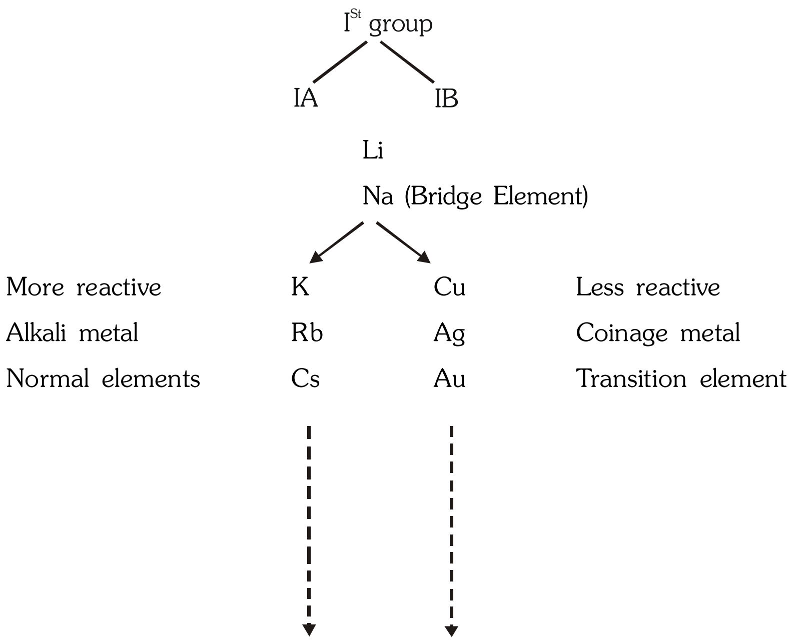
$$\begin{pmatrix} \text{Co} & \text{Ni} \\ 58.9 & 58.6 \end{pmatrix} \qquad \begin{pmatrix} \text{Th} & \text{Pa} \\ 232 & 231 \end{pmatrix}$$

(d) Like elements were placed in different groups :

There were some elements like Platinum (Pt) and Gold (Au) which have similar properties but were placed in different groups in Mendeleev's table.

Pt	Au
VIII	IB

(e) Unlike elements were placed in same group :



Cu, Ag and Au placed in Ist group along with Na, K etc. While they differ in their properties (Only similar in having ns^1 electronic configuration)

(f) It was not clear that **'lanthanides** and **Actinides'** are related with IIIA group or IIIB group.

(g) Cause of periodicity : Why physical & chemical properties repeated in a group.

MODERN PERIODIC TABLE (MODIFIED MENDELEEV PERIODIC TABLE) :

(i) It was proposed by Moseley.

(ii) Modern periodic table is based on atomic number.

(iii) Moseley did an experiment in which he bombarded high speed electron on different metal surfaces and obtained X-rays.

He found out that $\sqrt{\nu} \propto Z$ where ν = frequency of X-rays, Z = atomic number.

From this experiment, Moseley concluded that the physical and chemical properties of the elements are periodic function of their atomic number. It means that when the elements are arranged in the increasing order of their atomic number elements having similar properties gets repeated after a regular interval. This is also known as 'Modern Periodic Law'.

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

(v) **Characteristics of modern periodic table :**

(a) 9 vertical columns called groups.

(b) I to VIII group + 0 group of inert gases.

(c) Inert gases were introduced in periodic table by Ramsay.

(d) 7 horizontal rows called periods.

LONG FORM / PRESENT FORM OF MODERN PERIODIC TABLE :

(It is also called as 'Bohr, Bury, Rang & Werner Periodic Table)

(i) It is based on the Bohr-Bury electronic configuration concept and atomic number.

(ii) This model is proposed by Rang & Werner

(iii) 7 periods and 18 groups

(iv) According to I. U. P. A. C. 18 vertical columns are named as 1st to 18th group.

(v) The co-relation between the groups in long form of periodic table and in modern form of periodic table are given below –

IA	IIA	IIIB	IVB	VB	VIB	VIIB	VIII			IB	IIB	IIIA	IVA	VA	VIA	VIIA	0
1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18

(vi) Elements belonging to same group have same number of electrons in the outermost shell so their properties are similar.

Description of periods

Period	n	Period Sub shell	No. of elements	Element	Name of Period
1.	1	1s	2	$_{1}H$ – $_{2}He$	Shortest
2.	2	2s, 2p	8	$_{3}Li$ – $_{10}Ne$	Short
3.	3	3s, 3p	8	$_{11}Na$ – $_{18}Ar$	Short
4.	4	4s, 3d, 4p	18	$_{19}K$ – $_{36}Kr$	Long
5.	5	5s, 4d, 5p	18	$_{37}Rb$ – $_{54}Xe$	Long
6.	6	6s, 4f, 5d, 6p	32	$_{55}Cs$ – $_{86}Rn$	Longest
7.	7	7s, 5f, 6d, 7p	32	$_{87}Fr$ – $_{118}Uuo$	Complete

CONCLUSION

Period number = outermost shell

Number of element in a period = Number of electrons in a period subshell

DESCRIPTION OF GROUPS :

1^{st}/IA/Alkali metals

H = $1s^1$
Li = $1s^2$, $2s^1$
Na = $1s^2$, $2s^2\,2p^6$, $3s^1$
K = $1s^2$, $2s^2\,2p^6$, $3s^2\,3p^6$, $4s^1$
General electronic configuration = ns^1
Number of valence shell e^- = 1

2^{nd}/IIA/Alkaline earth metals

Be = $1s^2$, $2s^2$
Mg = $1s^2$, $2s^2\,2p^6$, $3s^2$
Ca = $1s^2$, $2s^2\,2p^6$, $3s^2\,3p^6$, $4s^2$
General electronic configuration = ns^2
(n = Number of shell)
Number of valence shell e^- = 2

13^{th}/IIIA/Boron Family

B = $1s^2$, $2s^2\,2p^1$
Al = $1s^2$, $2s^2\,2p^6$, $3s^2\,3p^1$
Ga = $1s^2$, $2s^2\,2p^6$, $3s^2\,3p^6\,3d^{10}$, $4s^2\,4p^1$
General electronic configuration = $ns^2\,np^1$
Number of valence shell e^- = 3

14^{th}/IVA/Carbon Family

C = $1s^2$, $2s^2\,2p^2$
Si = $1s^2$, $2s^2\,2p^6$, $3s^2\,3p^2$
Ge = $1s^2$, $2s^2\,2p^6$, $3s^2\,3p^6\,3d^{10}$, $4s^2\,4p^2$
General electronic configuration = $ns^2\,np^2$
Number of valence e^- = 4

15^{th}/VA/Nitrogen family/Pnicogen

(Used in fertilizer as urea)
N = $1s^2$, $2s^2\,2p^3$
P = $1s^2$, $2s^2\,2p^6$, $3s^2\,3p^3$
As = $1s^2$, $2s^2\,2p^6$, $3s^2\,3p^6\,3d^{10}$, $4s^2\,4p^3$
General electronic configuration = $ns^2\,np^3$
Number of valence shell e^- = 5

16^{th}/VIA/Oxygen family/Chalcogen

(Ore forming)
O = $1s^2$, $2s^2\,2p^4$
S = $1s^2$, $2s^2\,2p^6$, $3s^2\,3p^4$
Se = $1s^2$, $2s^2\,2p^6$, $3s^2\,3p^6\,3d^{10}$, $4s^2\,4p^4$
General electronic configuration = $ns^2\,np^4$
Number of valence shell e^- = 6

17^{th}/VIIA/Fluorine family/Halogens

(Salt forming)
F = $1s^2$, $2s^2\,2p^5$
Cl = $1s^2$, $2s^2\,2p^6$, $3s^2\,3p^5$
Br = $1s^2$, $2s^2\,2p^6$, $3s^2\,3p^6\,3d^{10}$, $4s^2\,4p^5$
General electronic configuration = $ns^2\,np^5$
Number of valence shell e^- = 7

18^{th}/Zero group/Inert gases / Noble gases

(Less reactive)
Ne = $1s^2$, $2s^2\,2p^6$
Kr = $1s^2$, $2s^2\,2p^6$, $3s^2\,3p^6\,3d^{10}$, $4s^2\,4p^6$
General electronic configuration = $ns^2\,np^6$ (except He)
Number of valence shell e^- = 8

(vii) 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
102	Un nil bium	Unb	Nobelium	No
103	Un nil trium	Unt	Lawrencium	Lr
104	Un nil quadium	Unq	Rutherfordium	Rf
105	Un nil pentium	Unp	Dubnium	Db
106	Un nil hexium	Unh	Seaborgium	Sg
107	Un nil septium	Uns	Bohrium	Bh
108	Un nil octium	Uno	Hassium	Hs
109	Un nil ennium	Une	Meitnerium	Mt
110	Un un nilium	Uun	Darmstadtium	Ds

(viii) Identification of group, period and block :

(A) When atomic number is given :

Step I : $71 \geq Z \geq 58 \Rightarrow$ Lanthanoids (6th Period)

$103 \geq Z \geq 90 \Rightarrow$ Actinoids (7th Period)] f-block

Group number = IIIB (largest group of periodic table)

Step II : Z = 104 to 118 (Period number = 7)

Group number = last two digits in atomic number of element

Example : Z = 104

Group no. = 4

Step III : Group number = 18 + given atomic number – atomic number of next noble gas

If the value of this formula is negative then use 32 instead of 18 in formula.

(B) When electronic configuration is given

Period number (n) = number of outermost shell/Highest shell number.

Block identification :

- If np electron present then p - block ($ns^2\ np^{1-6}$)

  group number = 12 + np electrons

- If np electron absent then s/f/d block

  If $(n-2)f^0\ (n-1)d^0\ ns^{1-2}$ = s block

  group number = ns electrons

  If $(n-2)f^{1-14}\ (n-1)d^{0-1}\ ns^2$ = f block

  group number = IIIB

- If any other configuration or $(n-1)d^{1-10}\ ns^{0-2}$ (d-block)

group number = (n–1)d electron + ns electron

Bohr's Classification

Inert gases	**Normal or representative elements**	**Transition element**	**Inner transition element**
outermost shell complete	outermost shell incomplete	n & n–1 shells incomplete either in atomic or ionic form	n, (n–1), (n–2) shells incomplete
6 element	s & p block element except inert gas 38 element	all d block element except = IIB (Zn, Cd, Hg & Uub) 36 element	f-block elements 28 elements

SOME IMPORTANT POINTS :

(a) 2nd period elements (Li, Be, B) Shows **diagonal relationship** with 3rd period elements (Mg, Al, Si). Because of same ionic potential value they shows similarity in properties.

(Ionic potential (ϕ) = $\frac{\textbf{Charge on cation}}{\textbf{Radius of cation}}$)

$$\begin{tikzcd} \text{Li} \arrow[dr] & \text{Be} \arrow[dr] & \text{B} \arrow[dr] & \\ \text{Na} & \text{Mg} & \text{Al} & \text{Si} \end{tikzcd}$$

(b) 3rd period elements (Na, Mg, Al, Si, P, S, Cl) except inert gases are called **typical elements** because they represent the properties of other element of their respective group.

(c) **TRANSURANIC ELEMENTS** :

Elements having atomic number more than 92 are known as transuranic element.
All transuranic elements are radioactive & artificial.

Th	Pa	U	Np —— Lr (f-block)	Unq —— Uub (d-block)
90	91	92	93 103	104 112

First man made element is Tc
First man made lanthanoid is Pm
All actinoids are radioactive but all lanthanoids are not artificial / man made (except Pm)

(d) The group containing most electro positive elements – GROUP IA.

(e) The group containing most electro negative elements – GROUP VIIA

(f) The group containing maximum number of gaseous elements–GROUP ZERO(18th)

(g) The group in which elements have generally ZERO valency – GROUP ZERO(18th)

(h) **In the periodic table**

Number of Gaseous elements – 11 (H, N, O, F, Cl + Noble gases)
Number of Liquid elements – 6 (Cs, Fr, Ga, Hg, Br, Uub)
Number of Liquid elements at room temp. – 2
Bromine is the only non-metal which exists in liquid form.
Number of Solid elements – 95 (if discovered elements are 112)

(i) 0/18 group have all the elements in gaseous form.

(j) 2nd period contains maximum number of gaseous elements. They are 4 **(N, O, F, Ne)**

(k) IIIB/3rd group is called longest group having 32 elements including 14 Lanthanides and 14 Actinides

Sc
Y
La...............Lanthanides (14)
Ac...............Actinides (14)

PERIODICITY

(A) In a period, the ultimate orbit remain same, but the number of e^- gradually increases.

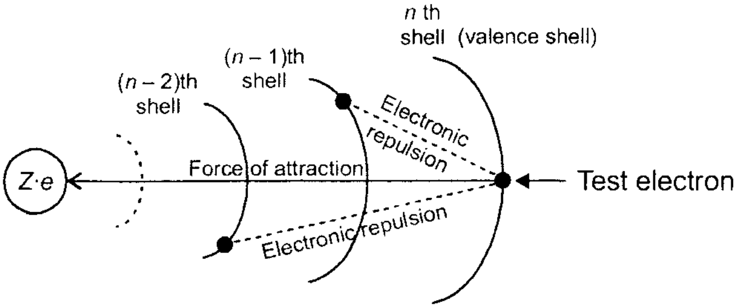
In a group, the number of e^- in the ultimate orbit remains same, but the values of n increases.

(B) Causes of periodicity :

(a) The cause of periodicity in properties is due to the same outermost shell electronic configuration coming at regular intervals.

(b) In the periodic table, elements with similar properties occur at intervals of 2, 8, 8, 18, 18 and 32. These numbers are called as magic numbers.

SCREENING EFFECT (σ) AND EFFECTIVE NUCLEAR CHARGE (Z_{eff}) :



(i) Valence shell e^- suffer force of attraction due to nucleus and force of repulsion due to inner shell electrons.

(ii) The decrease in force of attraction on valence e^- due to inner shell e^- is called screening effect or shielding effect.(i.e. total repulsive force is called shielding effect.)

(iii) Due to screening effect valence shell e^- experiences less force of attraction exerted by nucleus.
i.e. total attraction force experieced by valence electrons represented by a number is Z_{eff}.

(iv) There is a reduction in nuclear charge due to screening effect. Reduced nuclear charge is called effective nuclear charge.

(v) If nuclear charge = Z, effective nuclear charge = Z_{eff} , σ (Sigma)= Screening constant or shielding constant.

So, $\boxed{Z_{eff} = (Z - \sigma)}$

- **Slater's rule to know screening constant (σ)**

(a) For single electron species $\sigma = 0$

(b) Screening effect (S.E.) for two e^- species 0.30

Ex. In He ($1s^2$)

Screening effect of one 1s e^-. where $\sigma = 0.30$

$\therefore$ Zeff = Z – σ = 2 – 0.30 = 1.7

(c) Screening effect of each ns and np (Outermost orbit) electrons is 0.35

(d) Screening effect of each (n – 1) penultimate orbit s, p, d electrons is 0.85

(e) Screening effect of each (n – 2) and below all the e^- present in s, p, d, f is 1.0

From top to bottom in a group Z_{eff} remain constant

Group	Element Z_{eff}	**Li** 1.30	**Na** 2.20	**K** 2.20	**Rb** 2.20	**Cs** 2.20	**Fr** 2.20
Period	Element Z_{eff}	**Be** 1.95	**B** 2.6	**C** 3.25	**N** 3.90	**O** 4.55	**F** 5.20

For same shell shielding effect has the order as s > p > d > f (due to penetration effect)

Z_{eff} for different ions of an element

$$\boxed{Z_{eff} \propto \frac{\text{positive charge}}{\text{negative charge}}}$$

(i) Z_{eff} for different ions of an element
(ii) Z_{eff} for isoelectronic species.

(i) Z_{eff} for different ions of an element

Ex. $N^+ > N > N^- = Z_{eff}$

(ii) Z_{eff} of isoelectronic species

Ex. $H^- < Li^+ < Be^{+2} < B^{+3}$ ($2e^-$ species)

$N^{-3} < O^{-2} < F^- < Na^+ < Mg^{+2}$ ($10e^-$ species)

ATOMIC RADIUS

The average distance of valence shell e^- from nucleus is called atomic radius. It is very difficult to measure the atomic radius because –

(i) The isolation of single atom is very difficult.

(ii) There is no well defined boundary for the atom. (The probability of finding the e^- is 0 only at infinity).

So, the more accurate definition of atomic radius is –

- Half the inter-nuclear distance(d) between two atoms in a homoatomic molecule is known as atomic radius.
- This inter-nuclear distance is also known as bond length. Inter-nuclear distance depends upon the type of bond by which two atoms combine.

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

(A) Covalent radius (B) Metallic radius (C) van der Waal's radius

(A) Covalent Radius

In homonuclear diatomic molecules covalent radius is the half of the internuclear distance between two singly bonded homo atoms.

Ex. If internuclear distance of A–A(A_2) molecule is (d_{A-A}) and covalent radius is r_A then

$$d_{A-A} = r_A + r_A \quad \text{or} \quad 2r_A$$

$$r_A = \frac{d_{A-A}}{2}$$

Ex. In Cl_2 molecule, internuclear distance is $1.98A^0$ so $r_{cl} = \frac{1.98}{2} = 0.99\,\text{Å}$

(B) Metallic/Crystal Radius

Half of the inter nuclear distance between two adjacent metallic atoms in crystalline lattice structure.

$$\text{Metallic radius} \propto \frac{1}{\text{Metallic bond strength}}$$

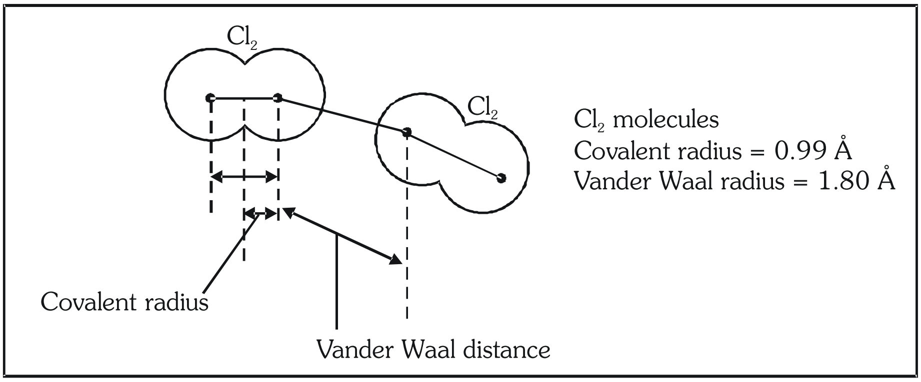
(C) Vander Waal's Radius

(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) Inert gas have only Vander Waal radius.

(d) In molecules of nonmetals solid both covalent and van der Waal's radius exists.



$$\boxed{\textbf{Vander Waal's radius} \cong \mathbf{2\times} \textbf{ covalent radius}}$$

$$\boxed{\textbf{Vander Waal's radius} > \textbf{ Metallic radius } > \textbf{Covalent radius}}$$

Ionic Radius

(i) Cationic Radius

(a) When an neutral atom loses e^- it converts into cation (+ve charged ion)

(b) Cationic radius is always smaller than atomic radius **because** after loosing e^- number of e^- reduces, but number of protons remains same, due to this Z_{eff} increases, hence electrons are pulled towards nucleus and atomic radius decreases, moreover after loosing all the electrons from the outer most shell, penultimate shell becomes ultimate shell which is nearer to nucleus so size decreases.

(c) Size of cation $\propto \dfrac{1}{\text{Magnitude of the charge or } Z_{eff}}$

Ex. (i) $Fe > Fe^{+2} > Fe^{+3}$

(ii) $Pb^{+2} > Pb^{+4}$

(iii) $Mn > Mn^{+2} > Mn^{+3} > Mn^{+4} > Mn^{+5} > Mn^{+6} > Mn^{+7}$

(ii) Anionic Radius

(a) When a neutral atom gains e^- it converts into anion [Negative charge ion]

(b) Anionic radius is always greater than atomic radius **because** in an anion e^- are more than protons and inter electronic repulsion increases, which also increases screening effect. So effective nuclear charge reduces, so distance between e^- and nucleus increases and size of anion also increases.

Ex. Flourine (Z=9)

	F	F^-
Proton	9	9
e^-	9	10

so $\dfrac{Z}{e} = \dfrac{9}{9} = 1$ $\dfrac{9}{10} = 0.9$ As Zeff of F^- is less than F so size of $F^- > F$

Size of isoelectronic species :

- Those species having same number of e^- but different nuclear charge forms isoelectronic series.
- For isoelectronic species the atomic radius increases with decrease in effective nuclear charge.

Species	**Ca^{+2}**	**K^{+}**	**Cl^{-}**	**S^{-2}**
Z	20	19	17	16
e	18	18	18	18
$\frac{Z}{e}$	$\frac{20}{18}$	$\frac{19}{18}$	$\frac{17}{18}$	$\frac{16}{18}$

Order of radius : $\underbrace{(S^{-2} > Cl^{-} > K^{+} > Ca^{+2})}_{18\ e^{-}}$, $\underbrace{(N^{3-} > O^{2-} > F^{-} > Na^{+} > Mg^{+2} > Al^{+3})}_{10\ e-}$

- **Factors affecting atomic size are :**

(a) **In a period**

$$\text{Atomic radius} \propto \frac{1}{Z_{eff}} \propto \frac{\text{negative charge}}{\text{positive charge}}$$

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

(b) **In a group**

Atomic radius ∝ number of shells

Li < Na < K < Rb < Cs

- **Periodic trends of atomic size :**

(i) **Across a period :** It decreases from left to right in a period as effective nuclear charge (Z_{eff})increases

Ex. Li > Be > B > C > N > O > F

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

Ex. Li < Na < K < Rb < Cs

Exceptions : Transition elements

$\underbrace{\text{Sc Ti V Cr Mn}}_{Z_{eff} > \text{ Screening effect}}$ $\underbrace{\text{Fe Co Ni}}_{Z_{eff} \approx \text{Screening effect}}$ $\underbrace{\text{Cu Zn}}_{Z_{eff} < \text{ Screening effect}}$

Order of covalent radius : Sc > Ti > V > Cr > Mn > Fe ≈ Co ≈ Ni < Cu < Zn

- **Lanthanide Contraction :**

(a) Outermost electronic configuration of inner transition elements is

$(n-2)\ f^{1-14}, (n-1)s^2p^6d^{0-1}, ns^2$ (n = 6 or 7)

(b) e^- enters in (n – 2) f orbitals

(c) Because of complicated structure of f-orbital and due to poor shielding f electrons, the outermost shell electrons get attracted towards nucleus.

(d) In I^{st}, 2^{nd} and 3^{rd} transition series, Radii– 3d < 4d ≈ 5d **(except III^{rd} B)**

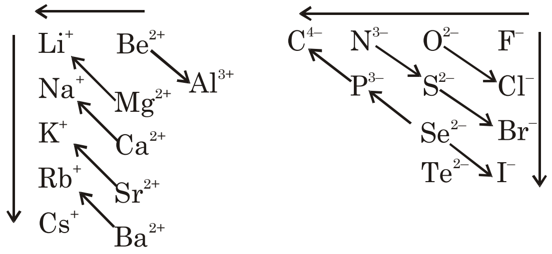
	IIIB	IVB	
size increases ↓	Sc	Ti ↓	size increases
	Y	Zr	Almost equal due to lanthanide contraction
	La	Hf	

- **Transition contraction :**

Note : While atomic size should increase down the group.

At. size of Ga ≈ At. size of Al, due to transition contraction. (Due to poor shielding of d electrons)

- **Special note: In the direction of arrow (→) ionic size increases.**



IONISATION POTENTIAL OR IONISATION ENERGY OR IONISATION ENTHALPY (IP / IE)

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

$$M_{(g)} + E_1 \longrightarrow M^+_{(g)} + e^-$$

- For any neutral atom ionization energy is always an endothermic process (ΔH = +ve)
- It is measured in eV/atom (electron volt/atom) or Kcal/mole or KJ/mole

Successive Ionization Energy

For an atom $M_{(g)}$ successive ionization energies are as follows -

$M_{(g)} + E_1 \longrightarrow M^+_{(g)} + e^-$ $\quad E_1$ = Ist Ionization energy

$M^+_{(g)} + E_2 \longrightarrow M^{+2}_{(g)} + e^-$ $\quad E_2$ = IInd Ionization energy

$M^{+2}_{(g)} + E_3 \longrightarrow M^{+3}_{(g)} + e^-$ $\quad E_3$ = IIIrd ionization energy

$\boxed{E_1 < E_2 < E_3 \ldots\ldots\ldots}$ (Always for an element)

FACTORS AFFECTING IONISATION ENERGY

In a period

(i) Effective nuclear charge (Z_{eff})

$$\boxed{\text{Ionisation Energy} \propto Z_{eff} \propto \frac{\text{positive charge}}{\text{negative charge}}}$$

Ion with high positive oxidation state will have high ionisation energy.

Ex. $Fe^{+3} > Fe^{+2} > Fe$

(ii) Stability of half filled and fully filled orbitals :

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

IE$_1$ N > O

$1s^2, 2s^2 2p^3$ $\quad$ $1s^2, 2s^2 2p^4$

After loosing one e^-, O attains electronic configuration of N, so IInd ionisation energy of O is more than N.

IE$_2$ N < O

$1s^2, 2s^2 2p^2$ $\quad$ $1s^2, 2s^2 2p^3$

(iii) Penetration power of sub shells

(a) Order of attraction of subshells towards nucleus (Penetration power) is s > p > d > f

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

Ex. **Be** **B**

$1s^2, 2s^2$ $\quad$ $1s^2, 2s^2 2p^1$

IE$_1$ Be > B

After loosing one e^-, B attains electronic configuration of Be, so IInd ionisation energy of B is more than Be.

IE_2 Be < B

$1s^2, 2s^1$ $1s^2, 2s^2$

In a group

Atomic size : $IE \propto \frac{1}{\text{atomic size}}$

COMPARISON OF IONISATION ENERGY

(i) **In a period :** Z_{eff} increases so removal of electron becomes difficult and ionisation energy increases.

Order of IE of 2nd period elements Li < B < Be < C < O < N < F < Ne

(ii) **In a group :** Size increase so ionisation energy decrease.

$$\xrightarrow{\text{Li} \quad \text{Na} \quad \text{K} \quad \text{Rb} \quad \text{Cs}}$$

Size increases, Ionisation Energy decreases

Exception :

- Ionisation Energy Ga > Al (due to Transition contraction)
- Ionisation Energy of 5d > 4d(due to lanthanide contraction)

  Ex. Hf > Zr

Application of ionisation energy :

(A) Metallic and non metallic character :

Generally for metals Ionisation Energy is low.

For Non-metals Ionisation Energy is high.

$$\boxed{\text{Metallic character} \propto \frac{1}{\text{IE}}}$$

(B) Reactivity of metals :

$$\boxed{\text{Reactivity of metals} \propto \frac{1}{\text{IE}}}$$

Reactivity of metals increases down the group as ionisation energy decreases.

(C) Stability of oxidation states of an element :

(a) If the difference between two successive ionisation energy of an element ≥ 16eV, then its lower oxidation state is stable.

Ex. $$\left.\begin{aligned} Na_{(g)} &\longrightarrow Na^{+}_{(g)} \\ Na^{+}_{(g)} &\longrightarrow Na^{2+}_{(g)} \end{aligned}\right] \Delta IE = 42.7eV$$

Difference between ionisation energy > 16 eV. So Na^+ is more stable.

(b) If the difference between two successive ionisation energy of an element ≤ 11 eV, then its higher oxidation state is stable.

Ex. $$\left.\begin{aligned} Mg_{(g)} &\longrightarrow Mg^{+}_{(g)} \\ Mg^{+}_{(g)} &\longrightarrow Mg^{2+}_{(g)} \end{aligned}\right] \Delta IE = 7.4eV$$

Difference of ionisation energy < 11 eV. So Mg^{+2} is more stable.

Ex. $$\left.\begin{aligned} Al_{(g)} &\longrightarrow Al^{+}_{(g)} \\ Al^{+}_{(g)} &\longrightarrow Al^{2+}_{(g)} \end{aligned}\right] \Delta IE = 12.8eV$$. So Al^+ is more stable

$$\left.\begin{aligned} Al^{+}_{(g)} &\longrightarrow Al^{2+}_{(g)} \\ Al^{2+}_{(g)} &\longrightarrow Al^{3+}_{(g)} \end{aligned}\right] \Delta IE = 6.0\ eV$$. So Al^{+3} is more stable.

Overall order of stability is $\boxed{Al^{+3} > Al^{+} > Al^{+2}}$

(D) To determine the number of valence electron of an element :

Number of valence electrons = number of lower values of IP before 1st highest jump.

ELECTRON AFFINITY/ELECTRON GAIN ENTHALPY (EA/ΔH_{eg})

The amount of energy released when an electron is added to the valence shell of an isolated gaseous atom known as Electron affinity.

$$X_{(g)} \quad + \quad e^- \quad \longrightarrow \quad X^-_{(g)} \quad + \quad \text{Electron Affinity}$$

Generally first electron addition of an isolated gaseous atom is an exothermic process
(except stable electronic configuration)

$$X_{(g)} \quad + \quad e^- \longrightarrow \quad X^-_{(g)} \quad + \quad \text{Electron Affinity}$$

ΔH_{eg_I} (first electron gain enthalpy) = – ve

Second electron addition of an isolated gaseous atom is always an endothermic process due to inter electronic repulsion.

$$X^-_{(g)} \quad + \quad e^- \quad \longrightarrow \quad X^{2-}_{(g)}$$

$\Delta H_{eg_{II}}$ (second electron gain enthalpy) = positive

Ex. $O_{(g)} \quad + \quad e^- \quad \longrightarrow O^-_{(g)} \quad ; \quad \Delta H_{eg_I} = -\ 141\ \text{KJ/mole}$

$O^-_{(g)} \quad + \quad e^- \quad \longrightarrow O^{-2}_{(g)} \quad ; \quad \Delta H_{eg_{II}} = +\ 744\ \text{KJ/mole}$

Net reaction $O_{(g)} \quad + \quad 2e^- \longrightarrow O^{-2}_{(g)} \qquad \Delta H_{eg_{II}} = +\ 603\ \text{KJ/mole}$

Formation of poly negative anion like O^{-2}, N^{-3}, C^{-4} etc. is always an endothermic process.

Electron affinity of neutral atom is equal to ionisation energy of its anion.

$$X_{(g)} + e^- \underset{\text{Ionisation Potential}}{\overset{\text{Electron Affinity}}{\rightleftarrows}} X^-_{(g)} \quad \left(\text{EA of } X_{(g)} = \text{IE of } X^-_{(g)}\right)$$

IE of neutral atom is equal to electron affinity of its cation

$$X_{(g)} \underset{\text{Electronic affinity}}{\overset{\text{Ionisation potential}}{\rightleftarrows}} X^+_{(g)} \qquad \left(\text{IE of } X_{(g)} = \text{EA of } X^+_{(g)}\right)$$

Factors affecting electron affinity :

(A) **Atomic size :** Electron Affinity $\propto \dfrac{1}{\text{Atomic size}}$

(B) **Effective nuclear charge (Z_{eff}) :** Electron Affinity $\propto Z_{eff} \propto \dfrac{\text{positive charge}}{\text{negative charge}}$

(C) **Stability of completely filled or half filled orbitals** : Electron affinity of elements having full-filled or half filled configuration is very less or zero so for these elements electron gain enthalpy $\left(\Delta H_{(eg)}\right)$ will be positive.

Ex. (i) $EA \Rightarrow ns^1 > ns^2$ (ii) $EA \Rightarrow np^2 > np^3$

Variation of electron affinity :

(i) **In 2nd period –**

$$\text{Li} \quad \boxed{\text{Be}} \quad \text{B} \quad \text{C} \quad \boxed{\text{N}} \quad \text{O} \quad \text{F} \quad \boxed{\text{Ne}}$$

$$\underbrace{\text{Ne} < \text{Be} < \text{N}}_{\text{endothermic}} < \text{B} < \text{Li} < \text{C} < \text{O} < \text{F}$$

(ii) **In Group :**

Electron affinity of 3rd period element is greater than electron affinity of 2nd period elements of the respective group.

F	Cl
[He] $2s^2 2p^5$	[Ne] $3s^2 3p^5$

Due to small size of fluorine, **electron density around the nucleus increases**. The incoming electron suffers more repulsion. In case of chlorine electron density decreases due to large size, decreasing order of electron affinity

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

Note : N & P have low electron affinity due to stable half filled configuration.

Electron affinity of the main-group elements (in electron volts)

1							18
H							**He**
+0.754	2	13	14	15	16	17	–0.5
Li	**Be**	**B**	**C**	**N**	**O**	**F**	**Ne**
+0.618	≤ 0	≤ +0.277	+1.263	–0.07	+1.461	+3.399	–1.2
Na	Mg	Al	Si	P	S	Cl	Ar
+0.548	≤ 0	+0.441	+1.385	+0.747	+2.077	+3.617	–1.0
K	**Ca**	**Ga**	**Ge**	**As**	**Se**	**Br**	**Kr**
+0.502	-	+0.03	+1.2	+0.81	+2.021	+3.365	–1.0
Rb	**Sr**	**In**	**Sn**	**Sb**	**Te**	**I**	**Xe**
+0.486	-	+0.3	+1.2	+1.07	+1.971	+3.059	–0.8

ELECTRONEGATIVITY (EN)

(i) The tendency of a covalently bonded atom to attract shared pair of electrons towards itself is called electronegativity.

(ii) A polar covalent bond of A – B may be broken as

$A — B \longrightarrow A^{\delta-} — B^{\delta+}$ (Electronegativity A > Electronegativity B)

depending on their tendency to attract bonded electron.

(iii) EN was explained by Pauling for the first time

Electronegativity of some other elements are as follows –

						H 2.1
Li 1.0	Be 1.5	B 2.0	C 2.5	N 3.0	O 3.5	F 4.0
Na 0.9	Mg 1.2	Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0
K 0.8						Br 2.8
Rb 0.8						I 2.5
Cs 0.7						
Fr 0.7						

In Pauling's scale, elements having almost same electronegativity are-

N = Cl = 3.0
C = S = I = 2.5
P = H = 2.1
Be= Al = 1.5
K = Rb = 0.8
Cs= Fr = 0.7

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

FACTORS AFFECTING ELECTRONEGATIVITY :

(A) **Atomic size**

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

Ex. F > Cl > Br > I

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

$$\boxed{\text{Electronegativity} \propto Z_{eff} \propto \frac{\text{positive charge}}{\text{negative charge}}}$$

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

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

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

--->

$Z_{eff} \uparrow EN\uparrow$

(C) **% s - character**

$$\boxed{\textbf{Electronegativity} \propto \textbf{\% s - Character}}$$

PERIODIC TRENDS OF ELECTRONEGATIVITY :

(A) Electronegativity decreases down the group.

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

APPLICATION OF ELECTRONEGATIVITY :

(A) **Metallic and non metallic nature :**

Generally metals have low electronegativity and non metals have high electronegativity, so we can say metallic character increases down the group but decreases along a period.

$$\boxed{\textbf{Non Metallic Nature} \propto \textbf{EN}}$$

(B) **Bond energy** : By increasing difference in electronegativity of bonded atoms, bond length decreases and hence bond energy increases

$$\boxed{\text{Bond energy} \propto \text{Electronegativity difference}}$$

(C) **Bond length (Schoemaker and Stevenson law) :**

If in a diatomic molecule electronegativities of A – B have more difference. Then actual bond length will be reduced. As per schoemaker & Stevenson– The reduction in bond length depends on the difference in electronegativities of atoms by following manner -

$d_{A-B} = r_A + r_B - 0.09\,(X_A - X_B)$

Here X_A is E.N. of A & X_B is E.N. of B

Ex. If bond length of F_2 = 1.44 Å, Bond length of H_2 = 0.74Å. Find out the bond length of H – F ? (EN of F is 4.0, EN of H is 2.1)

Solution. $d_{H-F} = r_F + r_H - 0.09\,(X_F - X_H)$

$\because\ r_F = 1.44 / 2 = 0.72\ \text{Å},\ r_H = 0.74/2 = 0.37\ \text{Å}$

$\therefore\ d_{H-F} = 0.72 + 0.37 - 0.09\,(4.0 - 2.1)$

$= 1.09 - (0.09 \times 1.9) = 1.09 - 0.171 = 0.919\ \text{Å}$

(D) **Acidic & Basic Strength :**

(i) Nature of hydroxides :

$$A{-}O{-}H{-}\begin{cases} \longrightarrow AO^{-} + H^{+} \text{ [Acidic]} \\ \longrightarrow A^{+} + OH^{-} \text{ [Basic]} \end{cases}$$

(a) If $X_A - X_0 \geq X_0 - X_H$ then AO bond will be more polar and will break up as

$A \text{——} OH \longrightarrow A^{+} + OH^{-}$ It shows basic nature

Ex. In NaOH

$X_O - X_{Na}\ (2.6) > X_O - X_H\ (1.4)$

(b) If $X_A - X_O \leq X_0 - X_H$ then OH bond will be more polar and will break up as

$A \text{—} O \text{——} H \longrightarrow H^{+} + AO^{-}$ It shows Acidic nature

In ClOH

$X_O - X_{Cl}\ (0.5) < X_O - X_H\ (1.4)$ So hydroxide is acidic

Order of acidic character :

$HNO_3 > H_2CO_3 > H_3BO_3$

$H_3PO_4 > H_3AsO_4 > H_3SbO_4$

$HOF > HOCl > HOBr > HOI$

$HClO_4 > HBrO_4 > HIO_4$

$H_2SO_3 > H_2SeO_3 > H_2TeO_3$

$HClO_4 > HClO_3 > HClO_2 > HClO$

(ii) Nature of oxides :

(a) Along a period acidic nature increases.

(b) Down the group basic nature increases

Li	Be	B	C	N	O	F
Na	Mg	Al	Si	P	S	Cl

$\longleftrightarrow$

Basic Amphoteric Acidic

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.

$\boxed{\text{Acidic strength of oxide and oxyacid} \propto \text{Electronegativity}}$

B_2O_3 CO_2 N_2O_5

$\longrightarrow$

EN increase, acidic nature increase.

$N_2O_5 > P_2O_5 > As_2O_5$

$N_2O_3 > P_2O_3 > As_2O_3 > Sb_2O_3$

$SO_3 > SeO_3 > TeO_3$

$\boxed{\text{Acidic nature} \propto \text{oxidation state}}$

Acidic properties increases with increasing oxidation state of an element

$N_2O_5 > N_2O_3$

$SO_3 > SO_2$

$Sb_2O_5 > Sb_2O_3$

(E) Nature of bonds :

According to Hanny & Smith formula

% ionic character = $16\ (X_A - X_B) + 3.5\ (X_A - X_B)^2$

Here X_A = Electronegativity of A

X_B = Electronegativity of B

If $X_A - X_B \geq 2.1$ Ionic % > 50% i.e. Ionic bond

If $X_A - X_B \leq 2.1$ Ionic % < 50% i.e. covalent bond

(F) Bond polarity $\propto \Delta$EN

$\Delta\chi = 0$ non-polar bond

$\Delta\chi \neq 0$ polar bond

Idientyfy the polar bond

C–Cl	$\Delta\chi \neq 0$
H–H	$\Delta\chi = 0$
O=O	$\Delta\chi = 0$
N–H	$\Delta\chi \neq 0$
Cl–F	$\Delta\chi \neq 0$

ELECTRONEGATIVITY SCALE :

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

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

where X_p is electronegativity on the basis of Pauling scale.

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

$$\boxed{X_p = \frac{\text{Ionisation Potential} + \text{Electron Affinity}}{5.6}}$$

$$\boxed{X_p = \frac{X_m}{2.8}}$$

- If ionisation potential and electron affinity are given in K.cal/mole then

$$\boxed{X_p = \frac{\text{Ionisation Potential} + \text{Electron Affinity}}{2 \times 62.5}}$$