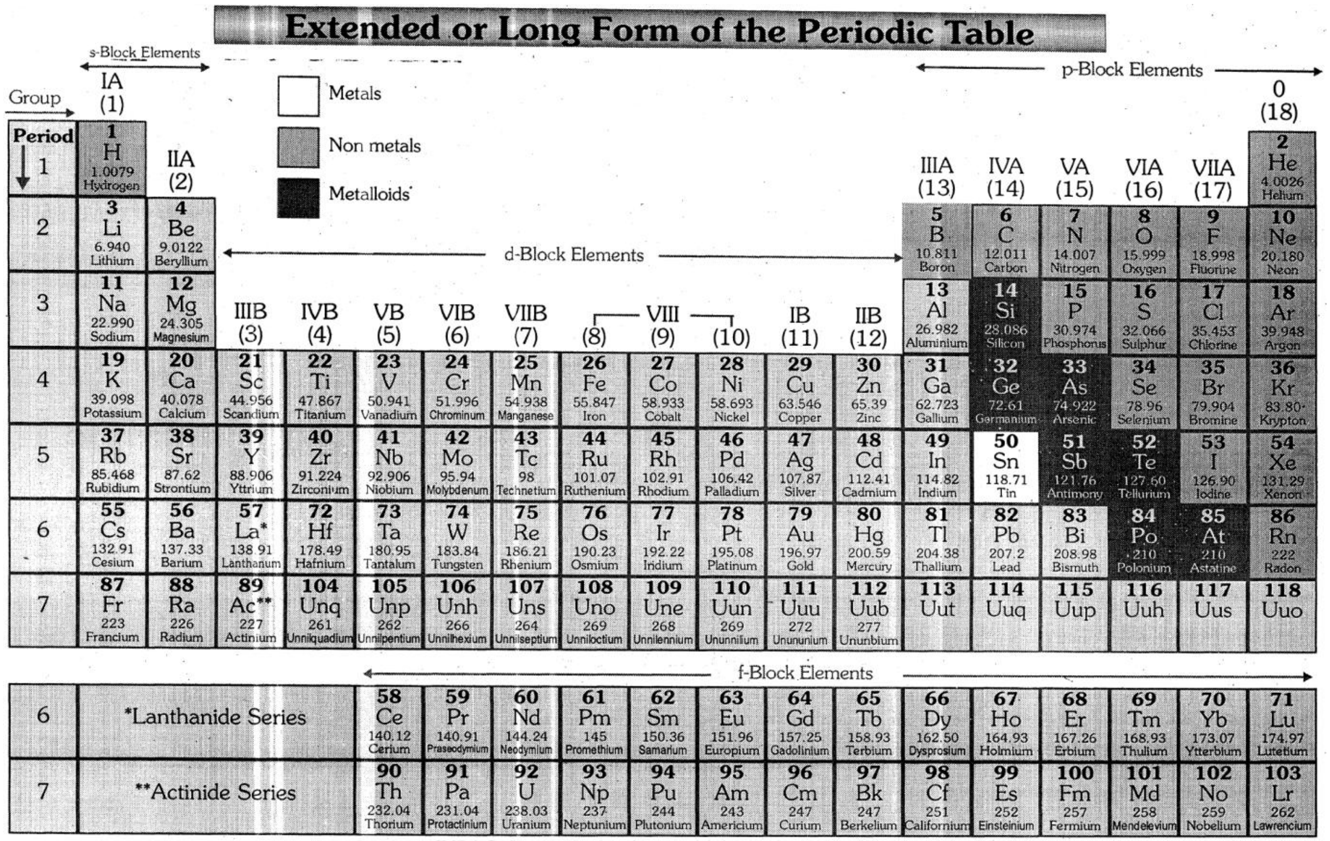


Extended or Long Form of the Periodic Table

☐ Metals
▨ Non metals
■ Metalloids

Period	IA (1)	IIA (2)	IIIB (3)	IVB (4)	VB (5)	VIB (6)	VIIB (7)	VIII (8)	VIII (9)	VIII (10)	IB (11)	IIB (12)	IIIA (13)	IVA (14)	VA (15)	VIA (16)	VIIA (17)	0 (18)
1	1 H 1.0079 Hydrogen																	2 He 4.0026 Helium
2	3 Li 6.940 Lithium	4 Be 9.0122 Beryllium											5 B 10.811 Boron	6 C 12.011 Carbon	7 N 14.007 Nitrogen	8 O 15.999 Oxygen	9 F 18.998 Fluorine	10 Ne 20.180 Neon
3	11 Na 22.990 Sodium	12 Mg 24.305 Magnesium											13 Al 26.982 Aluminium	14 Si 28.086 Silicon	15 P 30.974 Phosphorus	16 S 32.066 Sulphur	17 Cl 35.453 Chlorine	18 Ar 39.948 Argon
4	19 K 39.098 Potassium	20 Ca 40.078 Calcium	21 Sc 44.956 Scandium	22 Ti 47.867 Titanium	23 V 50.941 Vanadium	24 Cr 51.996 Chrominum	25 Mn 54.938 Manganese	26 Fe 55.847 Iron	27 Co 58.933 Cobalt	28 Ni 58.693 Nickel	29 Cu 63.546 Copper	30 Zn 65.39 Zinc	31 Ga 62.723 Gallium	32 Ge 72.61 Germanium	33 As 74.922 Arsenic	34 Se 78.96 Selenium	35 Br 79.904 Bromine	36 Kr 83.80 Krypton
5	37 Rb 85.468 Rubidium	38 Sr 87.62 Strontium	39 Y 88.906 Yttrium	40 Zr 91.224 Zirconium	41 Nb 92.906 Niobium	42 Mo 95.94 Molybdenum	43 Tc 98 Technetium	44 Ru 101.07 Ruthenium	45 Rh 102.91 Rhodium	46 Pd 106.42 Palladium	47 Ag 107.87 Silver	48 Cd 112.41 Cadmium	49 In 114.82 Indium	50 Sn 118.71 Tin	51 Sb 121.76 Antimony	52 Te 127.60 Tellurium	53 I 126.90 Iodine	54 Xe 131.29 Xenon
6	55 Cs 132.91 Cesium	56 Ba 137.33 Barium	57 La* 138.91 Lanthanum	72 Hf 178.49 Hafnium	73 Ta 180.95 Tantalum	74 W 183.84 Tungsten	75 Re 186.21 Rhenium	76 Os 190.23 Osmium	77 Ir 192.22 Iridium	78 Pt 195.08 Platinum	79 Au 196.97 Gold	80 Hg 200.59 Mercury	81 Tl 204.38 Thallium	82 Pb 207.2 Lead	83 Bi 208.98 Bismuth	84 Po 210 Polonium	85 At 210 Astatine	86 Rn 222 Radon
7	87 Fr 223 Francium	88 Ra 226 Radium	89 Ac** 227 Actinium	104 Unq 261 Unniquadium	105 Unp 262 Unnilpentium	106 Unh 266 Unnilhexium	107 Uns 264 Unnilseptium	108 Uno 269 Unniloctium	109 Une 268 Unnilennium	110 Uun 269 Ununnilium	111 Uuu 272 Unununium	112 Uub 277 Ununbium	113 Uut	114 Uuq	115 Uup	116 Uuh	117 Uus	118 Uuo

6	*Lanthanide Series	58 Ce 140.12 Cerium	59 Pr 140.91 Praseodymium	60 Nd 144.24 Neodymium	61 Pm 145 Promethium	62 Sm 150.36 Samarium	63 Eu 151.96 Europium	64 Gd 157.25 Gadolinium	65 Tb 158.93 Terbium	66 Dy 162.50 Dysprosium	67 Ho 164.93 Holmium	68 Er 167.26 Erbium	69 Tm 168.93 Thulium	70 Yb 173.07 Ytterbium	71 Lu 174.97 Lutetium
7	**Actinide Series	90 Th 232.04 Thorium	91 Pa 231.04 Protactinium	92 U 238.03 Uranium	93 Np 237 Neptunium	94 Pu 244 Plutonium	95 Am 243 Americium	96 Cm 247 Curium	97 Bk 247 Berkelium	98 Cf 251 Californium	99 Es 252 Einsteinium	100 Fm 257 Fermium	101 Md 258 Mendelevium	102 No 259 Nobelium	103 Lr 262 Lawrencium

IUPAC designations of groups of elements are given in brackets

PERIODIC TABLE

1.0 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 become 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) PROUT1S 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)

Atomic weight if H = 1

When n = number of hydrogen atom = 1, 2, 3,…..

Drawback or Limitation :

(i) Every element cannot be formed by Hydrogen.

(ii) Atomic weight of all element were no 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
1 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 observe that 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.

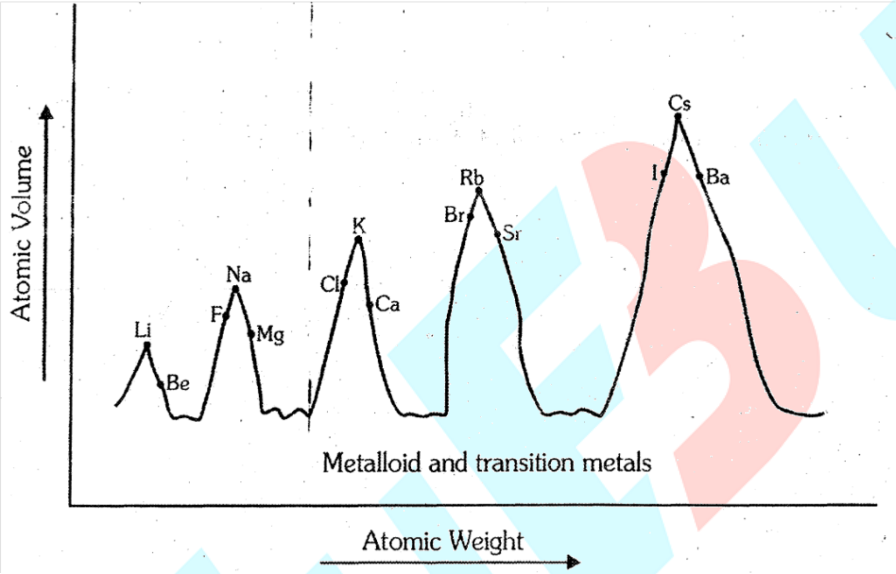
(c) Metalloids (Si, Ge, As, Sb, Te, Po, At) and transition metals occupy bottom part of the curve.

(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) MEDELEEV’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 ~re 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 element discovered after Mendeleev's periodic table were.

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

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

Atomic weight = Valency × 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 arid 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 2×4.5 = 9 and there was a space between U and B for this element in Mendeleev table.

Corrections were done ih 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.

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

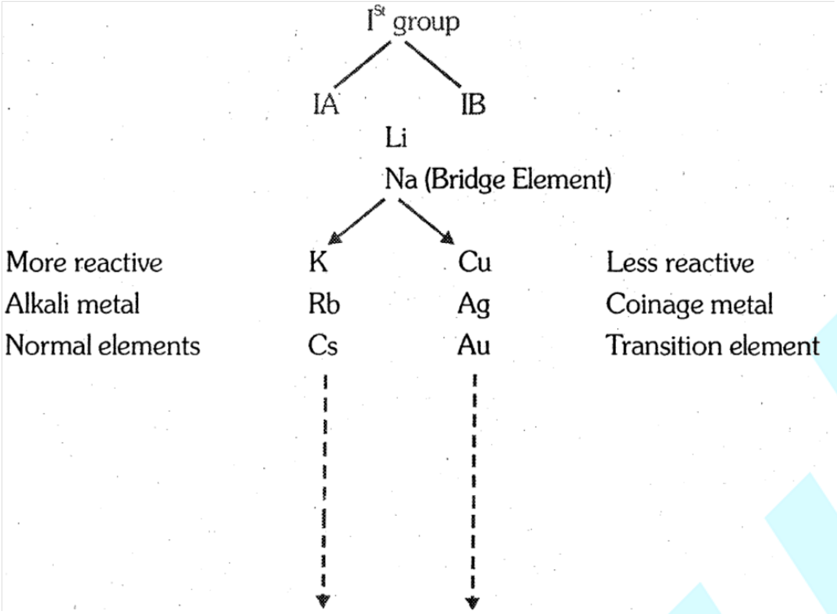
$$\begin{pmatrix} \text{Co} & \text{Ni} \\ 58.9 & 58.6 \end{pmatrix} \qquad \begin{pmatrix} \text{Th} & \text{Pa} \\ 232 & 231 \end{pmatrix}$$

(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 1st group along with Na, K etc. While they differ in their properties (Only similar in having ns1 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.

BEGINNER'S BOX-1

1. Mendeleev's periodic law is based on -
(1) Atomic number (2) Atomic weight
(3) Number of neutrons (4) None of the above

2. The first attempt to classify elements systematically was made by -
(1) Mendeleev (2) Newland
(3) Lather Meyer (4) Dobereiner

3. Atomic weight of an element X is 39, and that of element Z is 132 atomic weight of their intermediate element Y, as per dobereiner triad, will be
(1) 88.5 (2) 93.0 (3) 171 (4) 85.5

4. Which of the following is not a dobereiner triad
(1) Li, Na, K (2) Mg, Ca, Sr (3) Cl, Br, I (4) S, Se, Te

5. The law of triads is applicable to
(1) C, N, O (2) H, O, N (3) Na, K, Rb (4) Cl, Br, I

6. The law of triads is not applicable on
(1) Cl, Br, I (2) Na, K, Rb (3) S, Se, Te (4) Ca, Sr, Ba

7. Which of the following set of elements obeys Newland's octave rule -
(1) Na, K, Rb (2) F, Cl, Br (3) Be, Mg, Ca (4) B, Al, Ga

8. For which of the pair Newland octave rule is not applicable -

(1) Li, Na (2) C, Si (3) Mg, Ca (4) Cl, Br

9. Which of the following element was present in Mendeleev's periodic table?

(1) Sc (2) Tc (3) Ge (4) None of these

10. Is Fe, Co, Ni are dobereiner triad?

1.1 MODERN PERIODIC TABLE (MODIFIED MENDELEEV PEWODIC 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 $\sqrt{v} \propto Z$ where v = frequency of X -rays, Z = atomic number.

From this experiment, Moseley concluded theyt 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) Modem periodic law : The physical & chemical properties of elements are the periodic function of their atomic number.

(v) Characteristics of modem 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 M()DERN PERIODIC TABLE:

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

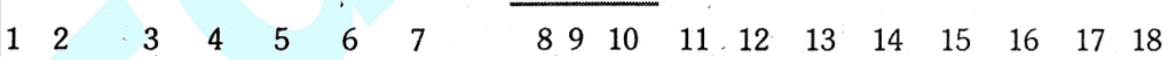
(i) It is based on the Bohr-Bury electronic configuration concept and atomic number.

(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 ling 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

1. Period number = outermost shell
2. Number of element in a period = Number of electrons in a period subshell

DESCRIPTION OF GROUPS :

1st/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

2nd/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

13th/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

14th/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

15th/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

16th/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	**18^{th}/Zero group/Inert gases / Noble gases**
(Salt forming)	(Less reactive)
F = $1s^2, 2s^2 2p^5$	Ne = $1s^2, 2s^2 2p^6$
Cl = $1s^2, 2s^2 2p^6, 3s^2 3p^5$	Kr = $1s^2, 2s^2 2p^6, 3s^2 3p^6 3d^{10}, 4s^2 4p^6$
Br = $1s^2, 2s^2 2p^6, 3s^2 3p^6 3d^{10}, 4s^2 4p^5$	General electronic configuration = $ns^2 np^6$ (except He)
General electronic configuration = $ns^2 np^5$	Number of valence shell e^- = 8
Number of valence shell e^- = 7	

(vii) NOMENCLATURE OF ELEMENTS :

(a) IPUAC 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, periods and block :

(A) **When atomic number is given :**

Step I: $\left.\begin{array}{l} 71 \geq Z \geq 58 \Rightarrow \text{Lanthanoids } (6^{th} \text{ Period}) \\ 103 \geq Z \geq 90 \Rightarrow \text{Actinoids } (7^{th} \text{ Period}) \end{array}\right]$ 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.

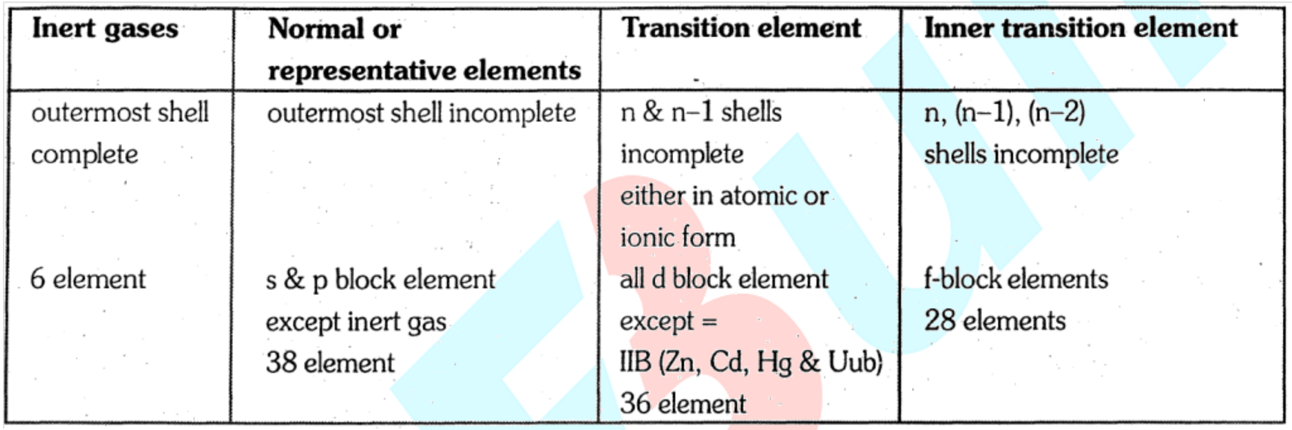
(B) **When electronic, configuration is given**

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

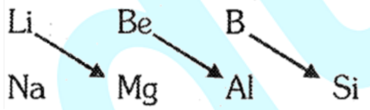


Inert gases	**Normal or representative elements**	**Transition element**	**Inner transition element**
outermost shell complete	outermost shell incomplete	n & n–1 shells incomplete either in atomic or ionic form	n, (n–1), (n–2) shells incomplete
6 element	s & p block element except inert gas 38 element	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) = $\dfrac{\text{Charge on cation}}{\text{Radius of cation}}$

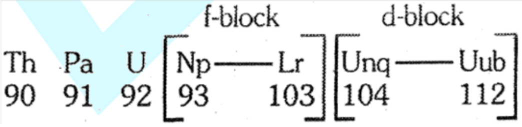


(b) 3rd period elements (Na, Mg, AI, Si, P, S, Cl) except inert gases are called typical elements because they represent the properties of other element of their respective group.

(c) TRANSURANIC ELEMENTS:

Elements having atomic number more than 92 are known as transuranic element.

All transuranic elements are radioactive & artificial.



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)

(b) The group containing most electro positive elements- GROUP IA.

(c) The group containing most electro negative elements - GROUP VIIA

(d) 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)
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)

BEGINNER'S BOX-2

1. Which of the following is best general electronic configuration of normal element.
(1) $ns^{1-2}\ np^{0-6}$ (2) $ns^{1-2}\ np^{1-5}$ (3) $ns^{1-2}\ np^{0-5}$ (4) $ns^{1-2}\ np^{1-6}$

2. Which of the following set of atomic numbers represents representative element
(1) 5, 13, 30, 53 (2) 11, 33, 58, 84 (3) 5, 17, 31, 54 (4) 9, 31, 53, 83

3. Which of the following electronic configuration does not belongs to same block as others:
(1) $[Xe]\ 4f^{14}\ 5d^{10}\ 6s^2$ (2) $[Kr]4d^{10}\ 5s^2$
(3) $[Kr]\ 5s^2$ (4) $(Ar)\ 3d^6\ 4s^2$

4. The electronic configuration of an element is $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^1$. What is the atomic number of next element of the same group which is recently discovered :-
(1) 20 (2) 119 (3) 111 (4) None

5. Which of the following electronic configurations in the outermost' shell is characteristic of alkali metals
(1) $(n-1)\ s^2p^6\ ns^2p^1$ (2) $(n-1)\ s^2p^6d^{10}\ ns^1$
(3) $(n-1)\ s^2p^6\ ns^1$ (4) $ns^2np^6\ (n-1)d^{10}$

6. Which of the following elements belong to alkali metals ?
(1) $1s^2,\ 2s^2\ 2p^2$ (2) $1s^2,\ 2s^2\ 2p^6,\ 3s^2\ 3p^6\ 3d^{10},\ 4s^2\ 4p^6,\ 5s^1$
(3) $1s^2,\ 2s^2\ 2p^5$ (4) None of these

7. Elements whose atoms have three outermost shells incomplete are called -
(1) s-block (2) p-block (3) d-block (4) f-block

8. Which of the following statement is wrong :-
(1) All the actinides are synthetic (man made) elements
(2) In the Lanthanides last electron enters in 4f orbitals
(3) Np_{93} onwards are transuranic elements
(4) Lanthanum is d-block element

9. Which of the following statement is wrong :-

(1) Total no. of liquid elements in the periodic table..... Six
(2) First metal element in the periodic table is.... Li
(3) All type of elements are present in 6th period u
(4) Iodine is a gaseous element.

10. An element 10 which is recently discovered is placed in 7th period and 10th group. IUPAC name of the element will be:-

(1) Unnilseptium (2) Ununnilium (3) Ununbium (4) None

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

- **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 $\sigma = 0.30$

$\therefore Z_{eff} = Z - \sigma = 2 - 0.30 = 1.7$

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

(d) Screening effect of (n–1) penultimate orbits, p, d electrons is 0.85

(e) Screening effect of (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	**Li**	**Na**	**K**	**Rb**	**Cs**	**Fr**
	Z_{eff}	1.30	2.20	2.20	2.20	2.20	2.20
Period	Element	**Be**	**B**	**C**	**N**	**O**	**F**
	Z_{eff}	1.95	2.6	3.25	3.90	4.55	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

$$Z_{eff} = \frac{\text{positive charge}}{\text{negative charge}}$$

(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)

1.3. ATOMICRADIUS

The average distance of valence shall 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 four categories -

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

(A) Covalent Radius

(SBCR - Single Bonded Covalent Radius).

(a) Covalent bonds are formed by overlapping of atomic orbitals.

(b) Internuclear distance is minimum in this case.

(c) 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 198Å so rd = $\frac{1.98}{2}$ = 0.99Å

(B) Ionic Radius

(i) Cationic Radius

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

(b) Cationic radius its 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 Z_{eff} of F^- is less than F so size of $F^- > F$

(c) 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	**K+**	**Ca^{+2}**	**S^{-2}**	**Cl^-**
Z	19	20	16	17
e	18	18	18	18
$\dfrac{Z}{e}$	$\dfrac{19}{18}$	$\dfrac{20}{18}$	$\dfrac{16}{18}$	$\dfrac{17}{18}$

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

(C) Metallic/Crystal Radius

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

(b) There is no overlapping of atomic orbitals, So **Metallic radius > Covalent radius**

(c) Metallic radius $\propto \dfrac{1}{\text{Metallic bond strength}}$

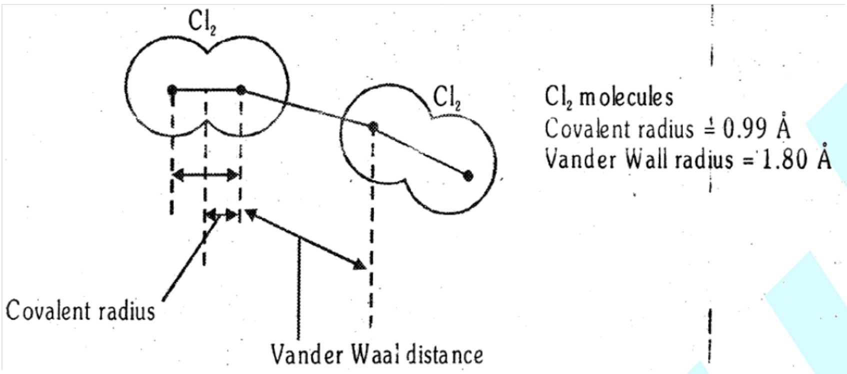
(D) Vander Waal's Radius

(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.



Vander Waal's radius ≅ 2 × covalent radius

Vander Waal's radius > Metallic radius > Covalent radius

- 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 variation of atomic size :**

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

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

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

Ex. Li < Na < K < Rb < Cs

Exceptions : Transition elements

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

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

- **Lanthanide Contraction :**

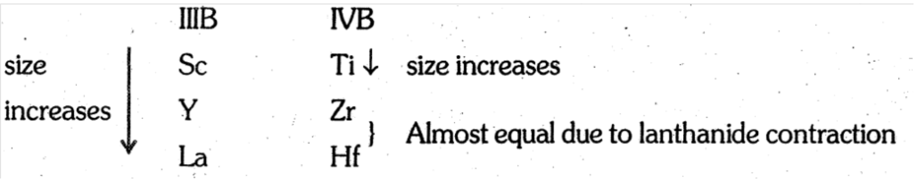
(a) Outermost electronic configuration of inner transition elements is

$(n-2)\ f^{1-14}, (n-1)s^2p^6d^{0-1}, ns^2$ (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 1st, 2nd and 3rd transition series, Radii-3d < 4d ≈ 5d (except III B)

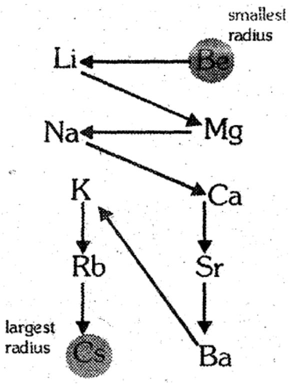


	IIIB	IVB	
size increases ↓	Sc	Ti ↓	size increases
	Y	Zr }	Almost equal due to lanthanide contraction
	La	Hf	

- Transition coatraction :

  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)
- **s-block size variation**



BEGINNER'S BOX-3

1. From the given set of species, point out the species from each set having least atomic radius:-

(1) O^{-2}, F^-, Na^+ (2) Ni, Cu, Zn

(3) Li, Be, Mg (4) He, Li^+, H^-

Correct answer is:-

(1) O^{-2}, Cu, Li, H^- (2) Na^+, Ni, Be, Li^+

(3) F^-, Zn, Mg, He (4) Na^+, Cu, Be, He

2. Which has the lowest anion to cation size ratio-

(1) LiF (2) NaF (3) CsI (4) CsF

3. Arrange the elements in increasing order of atomic radius Na, Rb, K, Mg :-

(1) Na, K, Mg, Rb (2) K, Na, Mg, Rb

(3) Mg, Na, K, Rb (4) Rb, K, Mg, Na

4. Which of the following pairs of elements have almost similar atomic radii :-

(1) Zr, Hf (2) Mo, W (3) Co, Ni (4) All

5. If the ionic radii of $K^{\oplus}$ and $F^{\ominus}$ are nearly the same (i.e. 1.34Å) then the atomic radii of K and F respectively are:-

(1) 1.34 Å, 1.34 Å (2) 0.72 Å, 1.96 Å

(3) 1.96 Å, 0.72 Å (4) 1.96 Å, 1.34 Å

6. For the element X, student mansi measured its radius as 102 nm, student Rohit as 203nm. And Ankur as 100 nm. using same apparatus. Their teacher explained that measurements were correct by saying that recorded values by three students were :-
(1) Crystal, van der Waal and Covalent radii
(2) Covalent, crystal and van der Waal radii
(3) van der Waal, ionic and covalent radii
(4) None is correct

7. Screening effect is not observed in :-
(1) He^+ (2) Li^{+2} (3) H (4) All of these

8. Arrange in orders of atomic and ionic radii :
(a) Ni, Cu, Zn (b) H^+, H, H^- (c) Ti, Zr, Hf (d) Be, Li, Na
(e) Cr, V, Ti, Sc (f) I^+, I, I^- (g) Sc, Y, La, Ac (h) Cl, Na, Rb
(i) Cu, Ag, Au (j) B, Be, Al, Mg (k) F, O, Cl, S

9. Which statement is false:-
(1) Screening effect increases down the group
(2) Z_{eff} increases down the group
(3) Z_{eff} increases in a period
(4) All of these

10. The screening effect of d- electrons is :
(1) Equal to the p- electrons (2) Much more than p - electrons
(3) Same as f- electrons (4) Less than p- electrons

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

(i) 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.

(Isolated → without any bonding with other atom)

(ii) Successive Ionization Energy

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

$M_{(g)} + E_1 \longrightarrow M^+_{(g)} + e^-$ E_1 = I Ionization energy
$M^+_{(g)} + E_2 \longrightarrow M^{+2}_{(g)} + e^-$ E_2 = II Ionization energy
$M^{+2}_{(g)} + E_3 \longrightarrow M^{+3}_{(g)} + e^-$ E_3 = III ionization energy

$E_1 < E_2 < E_3$……. Always for an element

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

(c) Ionization Energy is always an endothermic process ($\Delta H = +ve$)

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

FACTORS AFFECTING IONISATION ENERGY

In a period

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

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

Ion with high positive oxidation state will have high ionization 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 > 0

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

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

IE_2 N < 0

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

(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

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

IE_1 Be > B

After loosing one e^-, B attains electronic configuration of Be, so III 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.

$$\xrightarrow[\text{size increases, Ionisation Energy decreases}]{\text{Li Na K Rb Cs}}$$

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.

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

(B) Reactivity of metals :

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

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

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

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

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

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

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

(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.

BEGINNER'S BOX-4

1. IP_1 and IP_2 of Mg are 178 and 348 K. cal mol^{-1}. The enthalpy required for the reaction $Mg \rightarrow Mg^{2+} + 2e^-$ is :-

(1) +170 K.cal (2) + 526 K.cal (3) – 170 K.cal (4) – 526 K.cal

2. The IP_1, IP_2, IP_3, IP_4 and IP_5 of an element are 7.1, 14.3, 34.5, 46.8, 162.2 eV respectively. The element is likely to be:-

(1) Na (2) Si (3) F (4) Ca

3. Which of the following element has 2nd IP < 1st IP

(1) Mg (2) Ne (3) C (4) None

4. In which of the following the energy change corresponds to first ionisation potential only :-

(1) $X_{(g)} \rightarrow X^{+}_{(g)} + e$ (2) $X_{2(g)} \rightarrow X^{+}_{(g)} + e$

(3) $X_{(s)} \rightarrow X^{+}_{(s)} + e$ (4) $X_{(aq)} \rightarrow X^{+}_{(aq)} + e$

5. In the given process which oxidation state is more stable.

$M_{(g)} \longrightarrow M_{(g)}^{+}$ $IE_1 = 7.9$ eV

$M_{(g)}^{+} \longrightarrow M_{(g)}^{+2}$ $IE_2 = 15.5$ eV

(1) M^+ (2) M^{+2} (3) Both (4)None

6. The electronic configuration of some neutral atoms are given below :-

(A) $1s^2\ 2s^1$ (B) $1s^2\ 2s^2\ 2p^3$ (C) $1s^2\ 2s^2\ 2p^5$ (D) $1s^2\ 2s^2\ 2p^6\ 3s^1$

In which of these electronic configuration would you expect to have highest :-

(i) IE_1 (ii) IE_2

(1) C, A (2) B, A (3) C, B (4) B, D

7. The correct order of decreasing second ionization energy of Li, Be, Ne, C, B

(1) Ne > B > Li > C > Be (2) Li > Ne > C > B > Be

(3) Ne > C > B > Be > Li (4) Li > Ne > B > C > Be

8. In which of the following element has highest value of ionisation energy-

(1) Ti (2) Zr (3) Hf (4) None of these

9. What is the correct order of ionisation energy :

(1) $K < Cu < Cu^+ < K^+$ (2) $K < Cu^+ < Cu < K^+$

(3) $Cu^+ < K < Cu < K^+$ (4) $K^+ < Cu^+ < Cu < K$

10. Match the column.

Column-I **Valence electronic configuration**	**Column-II** **Successive ionisation energies**
(a) ns^1	(p) 19, 27, 36, 48, 270
(b) ns^2	(q) 16, 28, 34, 260
(c) $ns^2\ np^1$	(r) 18, 26, 230, 250
(d) $ns^2\ np^2$	(s) 14, 200, 220, 240

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

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

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

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

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

Δ_{egI} (first electron gain enthalpy) = –ve

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

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

ΔH_{egII} (second electron gain enthalpy) = positive

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

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

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

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

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

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

(6) IE of neutral atom is equal to electron affinity of its cation

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

(7) Factore 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 ($\Delta H_{(eg)}$) will be positive.

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

(8) Variation of electron affinity:

(i) In 2nd period – Li $\boxed{\text{Be}}$ B C $\boxed{\text{N}}$ O F $\boxed{\text{Ne}}$

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

(ii) In Group :

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

F	Cl
$[He]2s^22p^5$	$[Ne]\ 3s^23p^5$

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.

BEGINNER'S BOX-5

1. The correct order of electron affinity is :-

(1) Be < B < C < N
(2) Be < N < 13 < C
(3) N < Be < C < B
(4) N < C< B < Be

2. In the formation of a chloride ion, from an isolated gaseous chlorine atom, 3.8 eV energy is released, which would be equal to:

(1) Electron affinity of Cl- (2) Ionisation potential of Cl
(3) Electronegativity of Cl (4) Ionisation potential of Cl^-

3. $O_{(g)} + 2e^- \rightarrow O^{2-}_{(g)}$ $\Delta Heg = 603$ KJ/mole. The positive value of ΔHeg is due to :-
(1) Energy is released to add on 1 e^- to O^{-1} (2) Energy is required to add on 1 e^- to O^{-1}
(3) Energy is needed to add on $1e^-$ to O (4) None of the above is correct

4. The electron affinity values for the halogens shows the following trend :-
(1) F < Cl > Br > I (2) F < Cl < Br < I
(3) F > Cl > Br > I (4) F < Cl > Br < I

5. The process requiring the absorption of energyis.
(1) $F \rightarrow F^-$ (2)$Cl \rightarrow Cl^-$ (3) $O \rightarrow O^{2-}$ (4) $H \rightarrow H^-$

6. Second electron affinity of an element is:
(1) Always exothermic (2) Endothermic for few elements
(3) Exothermic for few elements (4) Always endothermic

7. Process, $Na^+_{(g)} \xrightarrow{I} Na_{(g)} \xrightarrow{II} Na_{(s)}$
(1) In (I) energy released, (II) energy absorbed
(2) In both (I) and (II) energy is absorbed
(3) In both (I) and (II) energy is released
(4) In (I) energy absorbed, (II) energy released

8. Which of the following configuration will have least electron affinity.
(1) ns^2np^5 (2) ns^2np^2 (3) ns^2np^3 (4) ns^2np^4

9. Which of the following will have the most negative electron gain enthalpy and which the least negative ?
(1) F, Cl (2) Cl, F (3) Cl, S (4) Cl, P

10. Which arrangement represents the correct order of electron gain enthalpy (with negative sign) of the given atomic species ?
(1) S < O < Cl < F (2) O < S < F < Cl (3) Cl < F < S < O (4) F < Cl < O < S

1.6 ELECTRONEGATIVITY (EN)

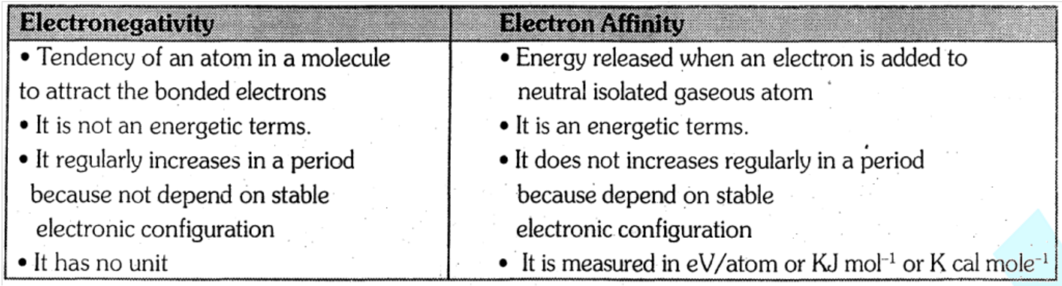
(i) The tendency of a covalently bonded atom to attract shared pair of electrons towards itself is called electronegativity.

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

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

depending on their tendency to attract bonded electron.

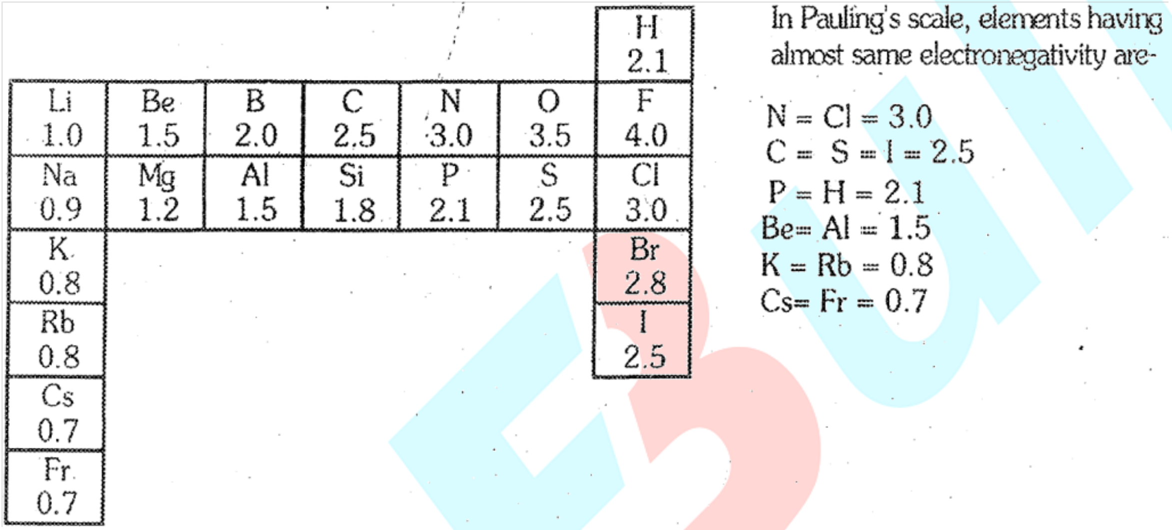
(iii) Difference between electronegativity and Electron Affinity :



Electronegativity	Electron Affinity
• Tendency of an atom in a molecule to attract the bonded electrons	• Energy released when an electron is added to neutral isolated gaseous atom
• It is not an energetic terms.	• It is an energetic terms.
• It regularly increases in a period because not depend on stable electronic configuration	• It does not increases regularly in a period because depend on stable electronic configuration
• It has no unit	• It is measured in eV/atom or KJ mol^{-1} or K cal mole^{-1}

(iv) 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						

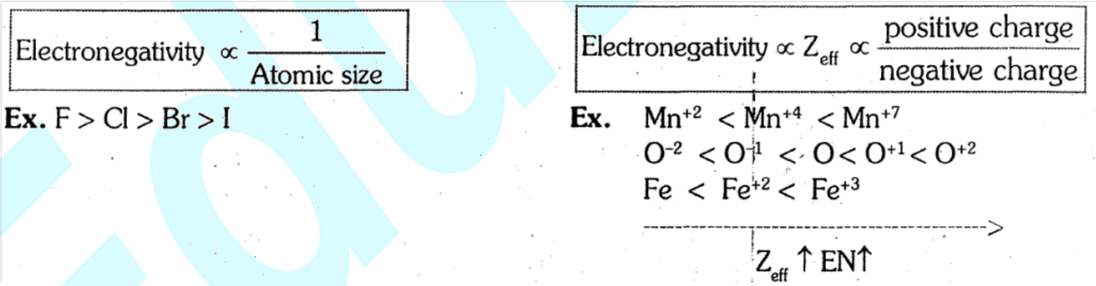
In Pauling's scale, elements having almost same electronegativity are-

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.

(v) FACTORS AFFECTING ELECTRONEGATMTY:

(A) Atomic size (B) Effective nuclear charge (Zeff)



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

Ex. F > Cl > Br > I

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

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

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

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

$\dashrightarrow$

$Z_{eff} \uparrow$ EN$\uparrow$

(C) % s – character

Electronegativity $\propto$ % s-character

(vi) PERIODIC TABLE & ELECTRONEGATMTY :

(A) Electro negativity decreases down the group.

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

(C) Electronegativity of Cs and Fr are equal, it is because from $_{55}$Cs to $_{87}$Fr only one shell increases but nuclear charge (Number of proton) increases by +32, so effect of nuclear charge palanced the effect of increase in number of shell

Electronegativity of F > Cl but Electron affinity of Cl > F

(D) In IIIA group, value of electronegativity is irregular when going down the group, because of transition contraction

Electronegativity of Ga > Electronegativity of Al

(vii) APPUCATIONOF 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.

[Non Metallic Nature ∝ EN]

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

[Bond energy ∝ Electronegativity difference]

HF > HCI > HBr > HI

(C) 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 H2 = 0.74Å. Find out the bond length of H–F?

(EN of F is 4.0, EN of H is 2.1)

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

Θ $r_F = 1.44/2 = 0.72$ Å, $r_H = 0.74/2 = 0.37$Å

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

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

(D) Acidic & Basic Strength:

(i) Nature of hydrides :

Stability of molecular ∝ Bond energy

Order of stability of hydrohalides : HF > HCl > HBr >Hl

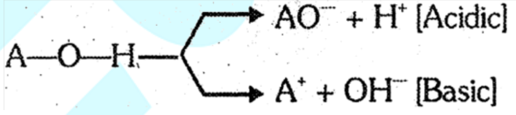
Order of acidic strength : HF < HCl < HBr < HI

In VA group : $NH_3 < PH_3 < AsH_3 < SbH_3 < BiH_3$

$$\xrightarrow{\hspace{10cm}}$$

Thermal stability decreases Acidic character increases

(ii) Nature of hydroxides :



(a) As per Gallis,

(i) In AOH if electronegativity of A is more than 1. 7 (Non metal) then it is acidic in nature.

(ii) If electronegativity of 'A' is less than 1.7 (metal) then AOH will be basic in nature

(b) If $X_A - X_O \geq X_O - X_A$ (X A = EN of A) then AO bond will be more polar and will break up as $A\text{–}OH \rightarrow A^+ + OH^-$ It shows basic nature

Ex. In NaOH

$X_O - X_{Na}$ (2.6) > $X_O - X_H$ (1.4) So hydroxide is basic

(c) If $X_A - X_{Na} \leq X_O - X_H$ then OH bond will be more polar and will break up as

$A - O - H \rightarrow H^+ + AO^-$ It shows Acidic nature

In ClOH

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

(iii) 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

i.e. when in periodic table te distance between the element ad oxygen increases, basic character increases.

$NO_2 > ZnO > K_2O$

$$\longrightarrow$$

Acidic character decreases

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.

B_2O_3 CO_2 N_2O_5

$$\longrightarrow$$

EN increase, acidic nature increase.

$HNO_3 > H_2CO_3 > H_3BO_3$ $HClO_4 > HBrO_4 > HIO_4$ $SO_3 > SeO_3 > TeO_3$

$H_3PO_4 > H_3AsO_4 > H_3SbO_4$ $N_2O_5 > P_2O_5 > As_2O_5$ $H_2SO_3 > H_2SeO_3 > H_2TeO_3$

HOF > HOCl > HOBr > HOI $N_2O_3 > P_2O_3 > As_2O_3 > Sb_2O_3$

[Acidic nature $\propto$ oxidation state]

(E) Nature of bonds :

(a) 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

(b) According to Gallis

$X_A - X_B \geq 1.7$ Ionic

$X_A - X_B \leq 1.7$ Covalent

If $X_A = X_B$; then A – B will be non polar. **Ex.** H–H, F–F

If $X_A > X_B$ and difference of electronegativities is small then

$A^{\delta-} - B^{\delta+}$ bond will be polar covalent

Ex. H_2O ($H^{\delta+} - O^{\delta-} - H^{\delta+}$)

If $X_A >> X_B$ and $X_A - X_B$ difference of electronegativities is high then

$A^- - B^+$ bond will be polar or ionic

(F) Nomenclature of inorganic compounds :

Prefix – less electronegative element

Suffix – More electronegative element

Ex. Cl_2O (Right) OCl_2 (Wrong)

In Dichloroxide the electronegativity of Cl is less than '0' i.e. why Cl is in prefix position.

OF_2 Oxygen difluoride

ICl Iodine chloride

(G) Bond polarity ∝ AEN

Ex. HF > HCl > HBr > HI

(viii) ELECTRONEGATIVITY SCALE:

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

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

where X_p is electronegativity on the basis of Pauling scale.

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

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

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

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

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

BEGINNER'S BOX-6

1. Which of the following is affected by stable configuration of an atom:-

(a) Electronegativity (b) Ionisation potential (c) Electron affinity

Correct answer is :-

(1) Only electronegativity

(2) Only ionisation potential

(3) Electron affinity and ionisation potential

(4) All of the above ·

2. Which of the following elements have the different value of electronegativity :-

(1) H (2) S (3) Te (4) P

3. Which is the correct ·order of electronegativity

(1) Cl > S > P > Si (2) Si > Al > Mg > Na

(3) F > Cl > Br > I (4) All

4. Electronegativity scale of pauling is based upon :-

(1) Bond length (2) Bond energy

(3) Atomic radius (4) Covalent radius

5. Correct order of electronegativity of N, P, C and I Si is :-

(1) N < P < C < Si (2) N > C > Si > P

(3) N = P > C = Si (4) N > C > P > Si

6. Outermost electronic configuration of the most electronegative element is :-

(1) ns^2np^3 (2) ns^2np^6 (3) ns^2 (4) ns^2np^5

7. Electro negativity of the following elements increases in the order.
(1) O, N, S, P (2) P, S, N, O (3) P, N, S, O (4) S, P, N, O

8. Give the correct order of electronegativity of central atom in following compounds -
(a)CH_3–CH_3 (b) CH_2=CH_2 (c) CH=CH
The correct order is -
(1) a > b > c (2) c > a > b (3) c > b > a (4) b > c > a

ANSWER KEY

BEGINNER'S BOX-1

1.	(2)	**2.**	(1)	**3.**	(4)	**4.**	(2)	**5.**	(4)
6.	(2)	**7.**	(3)	**8.**	(4)	**9.**	(4)	**10.**	(No)

BEGINNER'S BOX-2

1.	(3)	**2.**	(4)	**3.**	(3)	**4.**	(3)	**5.**	(3)
6.	(2)	**7.**	(4)	**8.**	(1)	**9.**	(4)	**10.**	(2)

BEGINNER'S BOX-3

1.	(2)	**2.**	(4)	**3.**	(3)	**4.**	(4)	**5.**	(3)
6.	(1)	**7.**	(1)	**8.**	( )	**9.**	(2)	**10.**	(4)

BEGINNER'S BOX-4

1.	(2)	**2.**	(2)	**3.**	(4)	**4.**	(1)	**5.**	(2)
6.	(1)	**7.**	(4)	**8.**	(3)	**9.**	(1)	**10.**	( )

BEGINNER'S BOX-5

1.	(2)	**2.**	(4)	**3.**	(2)	**4.**	(1)	**5.**	(3)
6.	(4)	**7.**	(3)	**8.**	(3)	**9.**	(4)	**10.**	(2)

BEGINNER'S BOX-6

1.	(3)	**2.**	(4)	**3.**	(4)	**4.**	(2)	**5.**	(4)
6.	(4)	**7.**	(2)	**8.**	(3)				