

INTRODUCTION :

The arrangement of all the known elements according to their properties in such a way that the elements of similar properties are grouped together in a tabular form is called 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 no. 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 p lace the metalloids.

(C) DOBEREINER TRIAD RULE :

- (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.**

$$\begin{bmatrix} Cl & Br & I\\ 35.5 & 80.0 & 127\\ x = \frac{35.5 + 127}{2} = 81.2 \\ \begin{bmatrix} Ca & Sr & Ba\\ 40 & 87.5 & 137\\ x = \frac{40 + 137}{2} = 88.5 \\ \end{bmatrix}$$

$$\begin{bmatrix} Ii & Na & K\\ 7 & 23 & 39\\ x = \frac{7 + 39}{2} = 23 \end{bmatrix}$$

Where x=average atomic weight

(iii) Other examples - (K, Rb, Cs), (P, As, Sb), (S, Se, Te), (H, F, Cl), (Sc, Y, La)

Key Points

- (a) For a dobereiner's triad, all the three elements should belong to the same group.
- (b) For a dobereiner's triad, Difference in the atomic number should be 8 or 18.

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 1stone. like in the case of musical vowels notation



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

Η

- 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, 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 difference of 18 elements instead of 8 element.
- (b) After the discovery of Inert gas and included in the period table it becomes the 8th element from Alkali metal so this law had to be dropped out.

(E) LOTHER 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 etc.) occupy the peak positions on the curve.
- (b) Less electropositive i.e. alkali earth metal (Be, Mg, Ca, Sr, Ba) occupy the descending position on the curve.

- (c) Metalloids (B, Si, As, Te, At etc.) 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 wt. and this become the base of Mendeleef's periodic table.



(F) TELLURIC HELIX: A three dimensional periodic table given by De-chan-chortois.

(G) MENDELEEF'S PERIODIC TABLE

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

(ii) Characteristic of Mendeleef'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 systamatic 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 groups in Mendeleev's Periodic table.

- (f) Each group upto VIIth is divided into A & B subgroups.'A' sub groups element are called normal elements and 'B' sub groups elements are called transition elements.
- (g) The VIIIth group was consists of 9 elements in three rows (Transitional metals group).
- (h) The elements belonging to same group exhibit similar properties.
- (iii) Merits or advantages of Mendeleef'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 mendeleef periodic table were.

Eka aluminium	-	Galium (Ga)
Eka Boron	-	Scandium (Sc)
Eka Silicon	-	Germanium (Ge)
Eka Mangense	-	Technetium (Tc)

(c) Correction of doubtful atomic weights : Correction were done in atomic weight of some 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 and there is no space in Mendeleev's table for this element. So, after correction, it was found that Be is actually divalent (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 Mendeleef'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 Mendeleef'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 didnot 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.

$$\begin{array}{c} \text{Ar} \quad \text{K} \\ 39.9 \quad 39.1 \end{array} \quad \left(\begin{array}{c} \text{Te} \quad \text{I} \\ 127.5 \quad 127 \end{array} \right) \quad \left(\begin{array}{c} \text{Co} \quad \text{Ni} \\ 58.9 \quad 58.6 \end{array} \right) \quad \left(\begin{array}{c} \text{Th} \quad \text{Pa} \\ 232 \quad 231 \end{array} \right)$$

(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 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.

	VIII		ron Cobalt Nickel Te Co Ni (Cu) 55.9 59 59 59	kuthenium Rhodium Palladium ku Rh Pd (Ag) .01.7 103.0 106.5		Semium Iridium Platinum Ds Ir Pt (Au) 91 193 194.9		RO4	
	IIA	Fluorine F 19.0 Chlorine Cl 35.45	Manganese I Mn Bromine E Bromine Br Br Br	- F F I lodine 126.9	1	-		R2O7 SOMPOUNDS RH	
SLN	VI	Oxygen O 16.00 Sulphur S 32.06	Chromium Cr 52.1 Selenium Se 79	Molybdenum Mo 96.0 Tellurium Te 127.6	1	Tungsten W 184	Uranium U 239	LINE OXIDES RO ₃ JS HYDROGEN C RH2	shad aanlian
UPS OF ELEME	Λ	Nitrogen N 14.04 Phosphorus P	Vanadium V 51.4 Arsenic As 75	Niobium Nb 94.0 Antimony Sb 120.0		Tantalum Ta 183 Bismuth Bi 208	1	HIGHER SAI R2O5 GHER GASEOU RH3	Table mible
GROI	N	Carbon C 12.0 Silicon Si 28.4	Titanium Ti 48.1 Germanium 72.3	Zirconium Zr 90.6 Tin Sn 119.0	Cerium Ce 140	- Lead Pb 206.9	Thorium Th 232	RO 2 HI RH 4	Daviadio
	Ш	- Boron B 11.0 Aluminium Al	Scandium Sc 44.1 Gallium Ga	Yttrium Y 89.0 Indium In 114.0	Lanthanum La 139	Ytterbium Yb 173 Thallium Tl 204.1	T	R203	oolob aa Moo
	п	- Beryllium Be 9.1 Magnesium Mg 24.3	Calcium Ca 40.1 Zinc Zn 65.4	Strontium Sr 87.6 Cadmium Cd 112.4	Barium Ba 137.4	- Mercury Hg 200.0	Radium Ra 224	RO	Fic
	I	Hydrogen H 1.008 Lithium Li 7.03 Sodium Na 23.5	Potassium K 39.1 Copper Cu 63.6	Rubidium Rb 85.4 Silver Ag 107.9	Caesium Cs 132.9	- Gold Au 197.2	ī	$\mathbb{R}_2\mathbb{O}$	
	0	- Helium He 4.0 Neon Ne 19.9	Argon Ar 38	Krypton Kr 81.8	Xenon Xe 128	¥.	ï	R	
SERIES		3 7 1	4 2	6	∞ o	10	12		

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(H) MODERN PERIODIC TABLE (MODIFIED MENDELEEF 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 $\sqrt{v} \propto Z$ where v = 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^{st} to VIII group + 0 group of inert gases.
- (c) Inert gases were introduced in periodic table by Ramsay.
- (d) 7 horizontal rows called periods.
- (I) 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 vertical columns (groups)
- (iv) According to I. U. P. A. C. 18 vertical columns are named as Ist 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

Noble gases	18	0	He 2	1.s ²	10	$Ne^{2s^2 2p^6}$	18	$\operatorname{Ar}_{3s^2 3p^6}$	36	Kr	45*4p°	40 ;	Xe seleti	98	μ'n	$6s^26p^6$		I			57	0		S^2			
	2		17	VII B	6	$F_{2s^2 2p^5}$	17	$\operatorname{Cl}_{3s^2 3p^5}$	35	Br	$4s^24p^5$	53	1 5,25.5	de se	×+	$6s^26p^5$		I		71	Lu	103	Lr	$^{2}5f^{14}6d^{1}7$			
elements	OMBE		16	VIB	8	$O_{2s^2 2p^4}$	16	$s^{2s^{23}p^{4}}$	34	Se	$4s^24p^4$	27	Te 5,25.4	<i>qc sc</i>	Do	$6s^{2}6p^{4}$		Uuh		70	Yb Adas de 2	102	No	$5f^{14}6d^{0}7s$	e outer	IUPAC	0 0 0 0
itative (15	VΒ	7	N $2s^2 2p^3$	15	${}^{\mathrm{P}}_{3s^23p^3}$	33	As	$4s^24p^3$	51	Sb 525.3	<i>dc.sc</i> 83	.id	$6s^26p^3$		I		69	$Tm_{\ell^{13} \in J^0 \mathcal{E}_{G^2}}$	101	Md	$f^{13}6d^07s^2$	nd state	1984 VIIB an	
Represer			14	IV B	9	$C_{2s^2 2p^2}$	14	$\operatorname{Si}_{3s^{2}3p^{2}}$	32	Ge	$4s^2 4p^2$	20 2	Sn 525.2	qc sc 82	40	$6s^26p^2$	114	Uuq		68	Er	100	Fm	r126d°7525	td grou	ith the	, mi
			13	III B	5	$^{\mathrm{B}}_{2s^{2}2p^{1}}$	13	$\operatorname{AI}_{3s^2 3p^1}$	31	Ga	$4s^24p^1$	49 ,	1n د 2ء - 1	de se	Ĺ.	$6s^{2}6p^{1}$		I		67	Ho	66	Es	$16d^{0}7s^{2}5f$	bers an	unce w	, 10H V
							- 5	II B	30	Zn	3d'45'	48	Cd	80 80	п	$5d^{10}6s^{2}$	112	Uub		99	Dy 5 dec2 de	98	Cf	6d ⁰ 7s ² 5f	uic num	accordc	VI ha
							11	IB	29	Cu tů	3d ⁴ 8 ¹	4 <	Ag A d0z J	79	ν.ν	$5d^{10}6s^{1}$	111	Uuu	lements	5	b]	14 on 41	3k	$d^{07S^2} 5f^{10}$	eir aton	-18 in	
							10		28	ïŻ	3d ⁶ 4s ⁴	40	Pd ^{010 L}	78	, ţ	$5d^{9}6s^{1}$	110	Ds	sition el	4 6	d I T	6 0	mB	$l^{1}T_{S}^{2}$ 5f ⁹ 6	with the	ered 1- mherine	ווחכווני
					nts	R	0	- IIIA	27	°C	3d ⁷ 4s ²	6	Rh A Me J	77	. t	$5d^{7}6s^{2}$	109	Mt	ner tran	79	G 4 2 6	6 17	0 7	$7s^{2} 5f^{7}6a$	nents 1	num blo	014 114
					elemer	UMBE	o	• •	26	Fe	3d ⁶ 4s ²	44	Ru Ade d	76	Č	5d ⁶ 6s ²	108	Hs	f-Inr	63	Eu	95	Am	$s^2 5f^76d^0$	he Elei	ps are	
		- 1	$\frac{1}{1S^1}$		ansition	OUP N	٢	VII A	25	Mn	3d ² 4s ²	4.	1c	75	Do	$5d^56s^2$	107	Bh		62	Sm Afe 206	94	Pu	5f6d ⁰ 7	tble of t	e grou	n i chim
					d-T ₁	– GR	y	VIA	24	C	3d ² 4s ¹	47	Mo 1,25c.h	74	/W	$5d^46s^2$	106	Sg		61	Pm	93	dN	5f6d17s	odic Ta	ns. Th	ווחותווח
							v	νA	23	, >`	3d ² 4s ⁴	41	Nb ¹ , ⁴ s, ¹	73	Ē	$5d^{3}6s^{2}$	105	Dþ		60	Nd Att dec2	92	Ŋ	$5f^36d^47s^2$	the Peri	iguratic	enit en
							r	ΥA	22	Ti	$3d^4s^4$	140	Zr	72	Пf	$4f^{14}5d^{2}6s^{2}$	104	Rf		59	Pr 1635 20622	91	Pa	$5f^26d^17s^2$	orm of	ic conf	ents.
							- ന	AIII	21	Sc	3d'4s'	39	Y $AA^{1}S_{c}^{2}$	57	* 1	5d ¹ 6s ²	89	Ac^{-1}		58	Ce	06	Th	r ⁰ 6d ² 7s ² 5	f guot	ectron	he elem
entative	JP CIUS	BER	7	ΠA	4	Be_{2s^2}	12	${ m Mg}_{3s^2}$	20	Ca	$4s^{2}$	38 2	Sr 2,2	56	р	Ба 6s ²	88	Ra_{7s^2}			4	Ŧ		51	g. 3.2	ωı	t
Repres	GROU	NUM	1	IA	3	Li 2s ¹	11	Na 3s ¹	19	К	4s ¹	37	Rb 5.1	55	ر	6s ¹	87	Fr_{7s^1}			$t^{0-1}6s^2$		ctinoids	ł" - 75'	Ы		
						7		ო BEK	IW	ОN 4	D	SIS	∾ bEF	[7	0	1			7	Lantl 4f"56	4	"A	5f″6ı			

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

Classification of Elements and Periodicity in Properties

(vii) Important Points :

(a) 2nd period elements (Li, Be, B) Shows diagonal relationship with 3rd period elements (Mg, Al, Si) so (Li, Be, B) are called Bridge elements. Because of same ionic potential value they shows similarity in properties. (Ionic potential = Charge/Radius)

- (b) 3rd period elements (Na, Mg, Al, Si, P, S, Cl) are called typical elements because they represent the properties of other element of their respective group.
- (c) In 6th period all types of elements are included (s, p, d & f)
- (d) No inert gas in 7th period.
- (e) Normal elements present in all periods.
- (f) Atomic number of last inert gas element is 86.
- (g) Long form of modern periodic table can be divided into four portions.
 - 1. Left portion IA & IIA s-block.
 - 2. Right portion IIIA to VII A + 0 group pblock.
 - 3. Middle portion IIIB to VIIB + VIII + IB IIB d-block.
 - 4. Bottom portion IIIB f-block elements
- (h) The group containing most electro positive elements GROUP IA.

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

- (j) The group containing maximum number of gaseous elements–GROUP ZERO(18^{th})
- (k) The group in which elements have generally ZERO valency-GROUP ZERO(18th)

(l) 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)

- (m) No p-block elements in 1^{st} and 7^{th} periods.
- (n) Ist period has all the elements in gaseous form (H, He)
- (o) 0/18 group have all the elements in gaseous form.
- (p) 2nd period contains maximum number of gaseous elements. They are 4 (N, O, F, Ne)
- (q) IIIB/3rd group is called longest group having 32 elements including 14 Lanthanides and 14 Actinides Sc
 - Y

La....Lanthanides (14)

Ac....Actinides (14)

(viii) 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	Hassnium	Hs
109	Un nil ennium	Une	Meitnerium	Mt
110	Un un nilium	Uun	Darmstadtium	Ds

(ix) Merits of long form of periodic table :

- (a) Position of isotopes Atomic number of isotopes are similar, so different isotopes can be placed at same place in periodic table.
- (b) (Ar K) (Co Ni) (Te I) are now in increasing order of atomic number.
- (c) Lanthanides and actinides are in IIIB group.
- (d) In modern periodic table diagonal line separates out metals, metalloids and non metals.
- (e) Elements of same group have same general formula of electronic configuration of outer most shell.
- (x) Demerits of long form of periodic table :
- (a) Position of hydrogen is still controversial.
- (b) 'He' is a inert gas but it has different electronic configuration than other inert gas elements.
- (c) Lanthanides and actinides are still not placed in main frame.
- (d) Isotopes have different physical properties but have same place in periodic table.

CLASSIFICATION OF ELEMENTS

- (A) Bohr classification : Elements can be divided in four parts on the basis of electronic configuration.
- (i) Inert gas elements :
- (a) The elements in which Ultimate orbit is completely filled are called inert gas elements.
- (b) General electronic configuration is ns^2p^6 (Except He = $1s^2$)
- (c) Because of most stable configuration, they are very less reactive. Hence known as noble gas or inert gas element.
- (d) These elements are present in '0' group or 18th group and 1 to 6th period of periodic table.
- (e) Number of inert gas elements are 6(One in each period upto 6th) He, Ne, Ar, Kr, Xe, Rn
- (ii) Normal or Representative elements :
- (a) The elements in which ultimate orbit is incomplete, while penultimate orbits are complete are called normal elements.



- (c) These elements lies in group IA to VIIA and period Ist to 7^{th}
- (d) Elements of 2^{nd} period known as bridge elements.
- (e) Elements of 3rd period (Na, Mg, Al, Si, P, S and Cl) are called typical elements These are 7 in numbers.

(iii) Transition elements :

Old concept (Given by Mendeleev's) : They are called transition elements because their properties are intermediate of s and p- block elements or they transit the properties of metals to non metals.

Note : According to this concept, Zn, Cd, Hg and Uub are transition elements.

New concept (Given by Bohr): The elements in which both ultimate (n) as well as penultimate shells (n-1) are incomplete either in atomic state or in some oxidation state are called transition elements.

Note : According to this concept, Zn, Cd, Hg and Uub are not transition elements because they do not have incomplete penultimate shell either in atomic state or in some oxidation state.

$$Zn = [Ar] 3d^{10} 4s^2$$
 and $Zn^{+2} = [Ar] 3d^{10}$

(a) group – IIIB to VIIB + VIII + IB and Periods – 4^{th} to 7^{th}

(b) Electronic configuration $(n-1)d^{1-10}ns^{1 \text{ or } 2}$

(c) Total number of d-block elements = 40

Total number of transition elements = 36 (Except Zn, Cd, Hg and Uub)

Note : All transition elements are d-block elements but all d-block elements are not transition elements.

(iv) Inner transition elements :

- (a) The Elements in which all the three shells that is ultimate (n) penultimate (n-1) and pre or antipenultimate (n-2) shell are incomplete are called inner transition elements.
- **Ex.** $_{58}$ Ce = [Xe] $6s^2$, $5d^1$, $4f^1$
- (b) Electronic configuration $(n-2)f^{1-14}(n-1)d^{0 \text{ or } 1} \text{ ns}^2$
- (c) These are 28 in number.
- (d) Group IIIB and Period 6^{th} & 7^{th}
- (e) Inner transition elements are divided into two series.
- (I) Lanthanide series/Rare earth elements / Lanthenones : (Ce₅₈ - Lu₇₁,14 elements)

The first element of this series is Cerium & not Lanthanum.

In these elements, last electron enters into 4f subshell

They are present in IIIB group and 6th period of the periodic table.

Promethium $(_{61}$ Pm) is the only lanthanide which is synthetic and radioactive in nature.

(II) Actinide series /Man made elements /Actinones $(Th_{90}-Lw_{103} 14 \text{ elements})$

The first element of this series is Thorium & not Actinium because last electron enters into 5f subshell.

They are present in IIIB group and 7th period of the periodic table.

All the actinides are radioactive in nature. First three elements (Th, Pa, U) are found in nature while others are synthetic in nature. After U_{92} i.e. from $_{93}Np$ onwards elements are called transuranic elements

They are heavier than uranium and they are derived from uranium by nuclear reactions.

- (B) Classification on the Basis of Sub shell in which last e⁻ enters :
- (i) s-block elements :
- (a) In these elements last electron enters in s subshell.
- (b) Group IA, IIA + 0 group (He only) and Period -1 to 7^{th}
- (c) Electronic configuration $-ns^{1-2}$
 - IA group n = 1 to 7 and IIA group n = 2 to 7

- (d) Total s block elements are (14) (including H and He)
- (e) Total s block metals are (12) (excluding H), H is non-metal

General characteristics of s-block elements

- ²²¹ s-block elements are soft metal, that is why these elements have low melting and boiling point.
- $_{2\overline{21}}$ Their oxides and hydroxides are basic in nature.
- $2^{1/21}$ They act as reducing agent.
- ²^[2] They have very low ionisation energy and highly electropositive so, they form ionic compound.
- ²^[2] Flame of these elements have the property of showing different colour.

(ii) p-block elements :

- (a) Last e^- enters in p sub shell
- (b) Group IIIA (13) to VIIA (17) + 0 group (18) (except He) and Period 2^{nd} to 6^{th}
- (c) Electronic configuration $ns^2 np^{1-6}$ and Total p - block elements -(30)

Characteristic of p-block elements

- ²^[2] p-block possesses all three kind of elements i.e. metals, non-metals and metalloids.
- ^{2/21} Oxides of non-metals are acidic in nature.
- ²¹They form covalent compounds mostly.
- ²^[2] They are oxidising in nature.

(iii) d-block elements :

- (a) Last e^- enters in (n-1)d subshell
- (b) Group IIIB VIIB, VIII, IB, IIB or group 3 to 12 (IUPAC) and Period – 4th to 7th
- (c) Electronic configuration $-(n-1)d^{1-10}ns^{1 \text{ or } 2}$
- (d) Total d block elements (40) Total transition elements – (36)
 - If 112 elements are included in periodic table
- (e) IIB elements (Zn, Cd, Hg, Uub) are d-block elements but not transition elements.

General characteristics of d-block elements

- ²^[2] They all are metals, which are very hard having high melting and boiling point.
- Elements of this block situated in between 's' and p-block elements. So the character of this block elements lie between s and p-block elements
- $^{2\overline{21}}$ They show variable oxidation state.
- **Ex.** $Mn = Mn^{+2}$, Mn^{+3} , Mn^{+4} , Mn^{+5} , Mn^{+6} , Mn^{+7}
 - ${}^{\scriptscriptstyle 2\overline{\scriptscriptstyle 21}}$ They form lonic and covalent bond both.
 - ²^[2] They are good conductor of heat and electricity and form complex compounds.
 - ²^[2] Metals, which have unpaired electrons show paramagnetism.
 - ²^[2] They form alloys and most of the elements act as catalyst.

(iv) f-block elements :

- (a) Differentiating e^- enters in (n-2) f subshell.
- (b) Group IIIB and Period 6^{th} and 7^{th}
- (c) From atomic number 58 71, 6th period ; Lanthanide series $4f^{1-14} 5d^{0 \text{ or } 1} 6s^2$ 90 - 103, 1 7th period ; Actinide series $5f^{1-14} 6d^{0 \text{ or } 1} 7s^2$
- (d) Total number of f block elements (28) and All the actinides are radioactive elements.
- (e) Transuranic actinides are man made elements. $(Np_{93} Lw_{103})$
- (f) Lanthanides are found rarely on earth so these are called rare earth metals.

General characteristics of f-block elements -

- $_{2/21}$ All the f-block elements are heavy metals.
- ${}^{\scriptscriptstyle 2\overline{\scriptscriptstyle 21}}$ It shows high melting and boiling point.
- ²^[2] The most common oxidation state of these elements is +3. **[RPMT 2004]**
- $2\overline{21}$ The elements of 5f series are radioactive in nature.

Predicting atomic number of the succesive member in a group of family

(i) Magic Numbers :

Knowing the atomic number of the first member of a group, we can write the atomic number of the subsequent elements by adding given magic number

Ex.	Group	1 2 3	4, 5, 6, 7, 8, 9 10, 11, 12,	13, 14, 15, 16, 17	18
		IA IIA IIIB		IIIA IV V VI VII	'O'gp.
	I +	2 –		i ▼	8
	I II+	8 8 🕇	I ♥	8	8
DD	III+	8 8		18	18
RIC	IV+	18 18 18	18	18	18
PE	V+	18 18 18	32	32	32
	VI ⁺	32 32 32	32	_	_
•	VII				

(ii) In group IA-Atomic number of H is 1 and atomic number of other element will be as follows -

Magic number $\underbrace{\frac{H_1 1+2=3}{2}}_{2}$ $\underbrace{\frac{\text{Li } 3+8=11}{8}}_{8}$ $\underbrace{\frac{\text{Na } 11+8=19}{8}}_{8}$ $\underbrace{\frac{\text{K } 19+18=37}{18}}_{18}$ $\underbrace{\frac{\text{Rb } 37+18=55}{18}}_{18}$ Cs

Determination of period, block and

group of an element

- (i) **Period number:** The period no. of the element can be predicted from the principal quantum no. (n) of the valence shell.
- (ii) Block number: Last electron enter in which orbital is knows as block no.
- (iii) Group number: It is predicted from the number of electrons in the valence shell and penultimate shell .
- Ex.

S.No.	Electronic	Period	Block	Prediction	Group	
	Configuration	number	number	of Group	number	
1.	$[Ar]4s^2 3d^{10}4p^6, 5s^1$	5	S	No. of ns electrons	1	
2.	[Kr] $5s^2$, $4d^{10} 5p^2$	5	р	ns ele. +10 +np ele.	2+10+2 =14	
3.	[Rn] $7s^2$, $6d^4 5f^{14}$	7	d	ns ele. $+(n-1)d$ ele.	2+4 =6	
4.	$[Xe] 6s^2, 5d^1, 4f^{12}$	6	f	—	3/III B	

PERIODICITY

- (i) The regular gradation in properties from top to bottom in a group and from left to right in a period is called periodicity in properties.
- (a) In a period, the ultimate orbit remain same, but the number of e⁻ gradually increases.
- (b) In a group, the number of e⁻ in the ultimate orbit remains same, but the values of n increases.
- (ii) 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.

Periodic Properties

Valency : It is defined as the combining capacity of the elements. The word valency is derived from an Italian word "Valentia" which means combining capacity.

Old concept : Given by : Frankland

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

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

IA	IIA	IIIA	IVA	VA	VIA	VIIA
Nał	H MgH ₂	AlH_3	SiH ₄	PH ₃	H_2S	H–Cl
NaC	Cl MgCl ₂	AlCl ₃	SiCl ₄	PCl ₃	SCl_2	C–Cl
Valency 1	2	3	4	3	2	1

Note :

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

Valency with respect to oxygen : Valency of O'=2It is defined as twice the number of oxygen atoms attached with a particular atom.

IA	IIA	IIIA	A IVA	A VA	VIA	VIIA	
Na	a ₂ O Mg	O Al	$_{2}O_{3}$ SiC	$P_2 P_2 O$	5 SO3	Cl_2O	7
Valency 1	2	3	4	5	6	7	

Note :

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

New concept : This concept is based on the electronic configuration. According to this concept valency for IA to IVA group elements is equal to number of valence shell e^- and from VA to zero group, it is -

[8–(number of	valence	e⁻)].		
	Valer	ncy=No	o. of valer	nce e-
	•			•
	IA	IIA	IIIA	IVA
	ns^1	ns ²	ns^2np^1	ns^2np^2
Valence shell e ⁻	1	2	3	4
Valency	1	2	3	4

Note : All the elements of a group have same valencies because they have same number of valence shell electrons.

Density :

- (i) Density increases down the goup because increase in atomic mass is much greater than volume.
- (ii) Factors affecting density are
 - (a) Inner shell configuration
 - (b) Atomic weight
 - (c) Packing capacity
 - (d) Chemical bond
- (iii) In period : Density first increases till maximum and then decreases. (s-block to d-block increases, dblock to p-block decreases)
- (iv) In group:
- Ex. In VIIA group F and Cl are gases (Low density) Br is liquid (density 3.19 g/cm⁻³) I is solid (density 4.94 g/cm⁻³)
- (vi) **Exceptions** The density of sodium and magnesium is greater than K and Ca respectively. This can be explained by inner shell configuration.

Density Na > K

In K, 3d⁰ orbital is vacant, 3rd orbit has the capacity to hold 18e⁻ but it has only 8e⁻. Due to empty orbitals atomic volumes increases and so density decreases

(vii) Li is the lightest metal known.

	Valency=(8–no.of valence	e e-)
•		I	+
VA	VIA	VII	0
ns ² np ³	ns ² np ⁴	ns ² np ⁵	ns ² np ²
5	6	7	8
3	2	1	0
(8-5) = 3			(8-8) = 0

(viii) Highest density of solid metal Ir (22.63 gm/cc) and Os (22.6 gm/cc) and liquid metal Hg(13.6 gm/cc)

(ix) Order of density of elements :

Li < K < Na < Rb < Cs $Ca < Mg < Be < Sr < Ba$ $He < Ar < Kr < Rn$	Sc < Y < La Ti < Zr << Hf Zn < Cd < Hg
Hg < Au < Os < Ir $Sc < Zn < Fe < Hg$	Cr < Fe < Cu Fe < Cu < Pb Fe < Hg < Au

Boiling points and melting points :

(i) It is a property of aggregate of atoms and not of a single atom.

i.e. why it is a molecular propertry.

- (ii) Melting point of a solid depend on –
 (a) Structure of solid
 (b) Chemical bond
 - (c) Bond energy
- (a) Structure of solid made up of (i) atomic solid (ii) molecular solid (iii) metallic solid.

Order of decreasing melting point is :

Atomic solid > Metallic solid > Molecular solid

(b) Chemical bond among particles in solid. Order of bond strength is – Covalent bond > Metallic bond > Vander waal bond

Covalent bond > Metallic bond > Vander waal bond

(c) Bond energy – Covalent solids like diamond, SiO_2 etc have only covalent bonds between atoms so their bond energy is higher than molecular or metallic solids.

(iii) **In period :** Along the period from left to right boiling point and melting point first increases then decreases.

Alkali metals – Crystal structure BCC (low boiling point & melting point)

Transition metals – FCC (High boiling point & melting point)

Inert gases Lowest boiling point & melting point (Vander waal force)

Transition elements Highest melting point

Metals W(Tungston) maximum melting point (3410°C) Hg(Mercury) lowest melting point (-38°C)

Non metals

Carbon (In the form of diamond) Highest melting point (3727°C)

Helium Lowest melting point (-270°C)

- (iv) In Group:
- (a) In **s-block** elements boiling point & melting point decreases down the group.

Li, Na (solid) \rightarrow Cs, Fr (liquid)

- (b) In **d-block** elements boiling point & melting point increases down the group (due to lanthanide contraction, zeff increases and hence bond energy increases)
- (c) In **p-block** elements
- (i) From IIIA IVA group boiling point & melting point decreases down the group and from VA to '0' group, boiling point & melting point increases down the group.
 (Atomic or molecular weight ∞ vander waal force)
- (ii) boiling point and melting point of monoatomic molecules are lesser than diatomic molecules.'0' group < Halogens
- (iii) Atomic solid non metals like B, C and Si has higher boiling point and melting point due to strong covalent bond.
- (iv) Boiling point & melting point of molecular solids are less because of weaker vander waal force among molecules

(v) Order of melting point :

Ex. I_{2} .

Li > Na > K > Cs	C > Si > Ge
Be > Ca > Sr > Ba > Mg	C > B > Si
$F_2 < Cl_2 < Br_2 < I_2$	B > Al > Ga

He < Ne < Ar < RnZn > Cd > HgCu > Au > Ag

Screening effect (σ) and effective nuclear charge (Z_{aff}):

- (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 experienced by valence e⁻ is called Zeff.)
- (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, Then effective nuclear charge = $Z - \sigma$ [Where σ (Sigma)= Screening constant] So, Zeff = $(Z - \sigma)$ Slater's rule to know screening constant (σ):
- (a) Screening effect (S.E.) of one e^- of the 1s is 0.30
- **Ex.** In He $(1s^2)$

Screening effect of one 1s e⁻. where $\sigma = 0.30$ \therefore Zeff = Z - $\sigma = 2 - 0.30 = 1.7$

- (b) Screening effect of ns and np (Outermost orbit) electron is 0.35
- (c) Screening effect of (n-1) penultimate orbit s, p, d electrons is 0.85
- (d) Screening effect of (n − 2) and below all the e⁻ present in s, p, d, f is 1.0

Lithium		$1s^22s^1 (Z=3)$	Potassium (Z = 19)	2,8,8,1 n-2 n-1 n
Zeff	=	Z-σ	Zeff	$= Z - \sigma$
Zeff on 2s ¹ electron	=	$3 - (2 \times 0.85)$	Zeff	$= 19 - (10 \times 1 + 8 \times .85)$
	=	3 - 1.70		=19 - 16.8
	=	1.30		=2.2
Note: Only two elect	ron p	present in $(n-1)$ shell affects Z.		
Sodium ($Z = 11$)		2, 8, 1 -2 n-1 n		
Zeff =	11	$-(2 \times 1 + 8 \times .85)$	= 11 - 8.8 = 2.2	

(Effective Nuclear charge of elements of second period)

Element	Electronic Configaration	Z	σ of ns & np electron	σ(n-1) orbital	Total Screeing	Effective nuclear
					Constant	charge
			(a)	(b)	(a + b)	Z*=Z-o
Li ³	$1s^2 2s^1$	3	—	$0.85 \times 2 = 1.70$	1.70	1.30
Be^4	$1s^2, 2s^2$	4	1× 0.35=0.35	0.85× 2=1.70	2.05	1.95
B^5	$1s^2, 2s^2, 2p^1$	5	2× 0.35=0.70	0.85× 2=1.70	2.40	2.60
C^6	$1s^2, 2s^2, 2p^1$	6	3× 0.35=1.05	0.85× 2=1.70	2.75	3.25
N^7	1s ² ,2s ² ,2p ³	7	4× 0.35=1.40	0.85× 2=1.70	3.10	3.90
O^8	$1s^2, 2s^2, 2p^4$	8	5× 0.35=1.75	0.85× 2=1.70	3.45	4.55
F ⁹	$1s^2, 2s^2, 2p^5$	9	6× 0.35=2.10	0.85× 2=1.70	3.80	5.20

Periodicity :

Ex.

- (i) From left to right in a period Zeff increases
- (a) For normal/Representative elements, In a period nuclear charge increases by+1, while orbit remains same.
- (b) In a period e⁻ enters in outermost shell and screening effect of outermost electron is 0.35. So by increasing 1 nuclear charge, Zeff increases by 0.65 (1 0.35) = 0.65.
- (c) That is why in a period Zeff increases by 0.65 and hence atomic size decreases considerably.
- (d) In transition series Z increase by + 1 but screening effect increases by 0.85 So Zeff is 0.15 (1- 0.85 = 0.15) [Because e⁻ enters in (n - 1)orbit which has the value of $\sigma = 0.85$]

In transition series Zeff increases by very less amount, i.e. by 0.15 from left to right and hence atomic size remains almost constant.

Element	Sc	Ti	V	Cr	Mn	Fe	Со	Ni	Cu	Zn
Zeff	3.00	3.15	3.30	2.95	3.60	3.75	3.90	4.05	3.70	4.35

(ii) From top to bottom in a group Zeff remain constant

Element	Li	Na	Κ	Rb	Cs	Fr
Zeff	1.30	2.20	2.20	2.20	2.20	2.20

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-
 - ^{2/21} Half the inter-nuclear distance(d) between two atoms in a homoatomic molecule is known as atomic radius.
 - ²²¹ 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) Vander waal 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_2}) and covalent radius is r_A then

 $d_{A-A} = r_A + r_A$ or $2r_A$ $r_A = \frac{d_{A-A}}{2}$ **Ex.** In Cl₂ molecule, internuclear distance is 1.98A⁰ so $r_{cl} = \frac{1.98}{2} = 0.99 \text{ Å}$

(d) SBCR of O, N and C etc. elements can be determined by taking H₂O₂, N₂H₄, C₂H₆ respectively.

Case I

For Heteroatomic molecule with less electronegativity difference.

For A-B molecule electronegativities of A & B are approximately equal Ex. C-I

Electronegativity of C & I are approx equal (2.5) internuclear distance of C - I is 2.13Å

 $d_{C-I} = r_{C} + r_{I}$, (r_c is 0.77 Å) ∴ r_I = 2.13 - 0.77 = 1.36Å Case II

Heteroatomic molecule with more electronegativity difference :

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

Solved Examples

Ex.1 If bond length of $F_2 = 1.44$ Å, Bond length of $H_2 = 0.74$ Å. Find out the bond length of H - F? (EN of F is 4.0, EN of H is 2.1)

Sol. $d_{H-F} = r_F + r_H - 0.09 (X_F - X_H)$ $\therefore d_{F-F} = 1.44/2 = 0.72 \text{ Å}, d_{H-H} = 0.72/2 = 0.37 \text{ Å}$ $\therefore d_{H-F} = 0.72 + 0.37 - 0.09 (4.0 - 2.1)$ $= 1.09 - (0.09 \times 1.9) = 1.09 - 0.171 = 0.919 \text{ Å}$

(B) 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 its Zeff increases, hence electrons pulls 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 $\propto \frac{1}{\text{Magnitude of the charg e or Zeff}}$

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

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

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

- (ii) Anionic radius
- (a) When neutral atom gains e^{-it} converts into anion

- (b) Anionic radius is always greater than atomic radius because In an anion e⁻ are more than protons so effective nuclear charge reduces, and inter electronic repulsion increases, which also increases screening effect. 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 $\frac{Z}{\rho} =$	$\frac{9}{9} = 1$	$\frac{9}{10} = 0.9$

As Zeff of F^- is less than F so size of $F^- > F$

- (c) Size of iso electronic species :
 - ²²¹ Those species having same number of e⁻ but different nuclear charge forms isoelectronic series.
 - ^{2/21} For isoelectronic species the atomic radius increases with decrease in nuclear charge

Species	K^+	Ca^{+2}	Ar	S^{-2}	Cl-
Z	19	20	18	16	17
e	18	18	18	18	18
Z	19	20	18	16	17
e	18	18	18	18	18

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

(C) Metallic radius

- (a) Half of the nuclear distance between two adjacent metallic atoms in crystalline lattice structure.
- (b) there is no overlapping of atomic orbitals so, Metallic radius > Covalent radius
- (c) Metallic radius $\propto \frac{1}{\text{Metallic bond strength}}$
- (d) More metallic radius \rightarrow loose crystal packing \rightarrow less bond strength

(body centred packing)

(e) Less metallic radius \rightarrow Tight crystal packing FCC \rightarrow High bond strength

(Hexagonal close packing)

(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) Vander Waal radius > Covalent radius.
- (d) Inert gas have only Vander Waal radius.
- (e) In molecules of nonmetals solid both covalent and Vander Waal radius exists.



Vander Waal radius = 2 × Covalent radius

- (f) Vander Waal radius > Metallic radius > Covalent radius
- (g) Vander Waal force of attraction ∞ Molecular weight or atomic weight (in inert gases)
- (h) In a period from left to right Vander Waal radius decreases.
- (i) In a group of non metal from top to bottom its values increases.

Factors affecting atomic size are :

(a) Atomic radius $\propto \frac{1}{\text{Effective nuclear charge (Zeff)}}$

(b) Atomic radius ∞ number of shells

Li < Na < K < Rb < Cs

- (c) Atomic radius \propto Screening effect
- (d) Atomic size \propto Magnitude of –ve charge $O < O^- < O^{-2}$

(e) Atomic radius
$$\propto \frac{1}{Magnitude \text{ of } + \text{ve charge}}$$

 $Mn > Mn^{+2} > Mn^{+3} > Mn^{+4}$

1

(f) Atomic radius $\propto \frac{1}{\text{Bond order}}$

N - N > N = N = N

Periodic variation of atomic size :

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

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

Ex. Li < Na < K < Rb < Cs

Exceptions : Transition elements

$$Sc > Ti > V > Cr Mn > Fe \approx Co \approx Ni < Cu < Zn$$

Lanthanoide Contraction :

(a) Outermost electronic configuration of inner transition elements is

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

- (b) e^{-} enters in (n-2) forbitals
- (c) Mutual screening effect of e⁻ is very less, because of complicated structure of f-orbital
- (d) Nuclear charge increases by one (+1) in lanthanoides and actinides so atomic size of these elements slightly decreases. It is known as lanthanoide contraction. Here Nuclear charge > Screening effect.
- (e) In Ist, 2^{nd} and 3^{rd} transition series, Radii- 3d $< 4d \approx 5d$ (except IIIrd B)

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

Transition contraction :

IIIA \rightarrow B<Al \approx Ga

Note : While atomic size should increases down the group.

- (a) At. size of Ga = At. size of Al, due to transition contraction.
- (b) In transition elements nuclear charge increases by 1.
- (c) but e^- enters in (n-1)d orbital exerts screening effect.
- (d) Screening effect of $(n-1)de^{-}$ balance the nuclear charge by 85%
- (e) Z_{eff} on increasing each electron = 1 0.85 = 0.15
- (f) Increase in nuclear charge is only 0.15 so atomic size remains almost constant.

Different orders of atomic and ionic radii :

$Na^+ < Ne < F^-$	Ni < Cu < Zn
Be < Li < Na	Cr < V < Ti < Sc
Cl < Na < Ar	Cu <au<ag< td=""></au<ag<>
$H_{+} < H < H_{-}$	$Ti < Zr \approx Hf$
$I^+ < I < I^-$	Sc < Y < La

IONISATION POTENTIAL OR IONIZATION ENERGY OR IONIZATION ENTHALPY

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

(Isolated \rightarrow Without any bonding with other atom)

- (ii) Successive Ionisation Energy :
- (a) For an atom M, successive ionisation energies are as follows -

$$M + E_1 \longrightarrow M^+ + e^-$$

$$E_1 = I^{st}$$
 Ionisation Potential

$$M^+ + E_2 \longrightarrow M^{+2} + e^-$$

 $E_2 = II^{nd}$ Ionisation Potential

 $M^{+2} + E_3 \longrightarrow M^{+3} + e^-$

 $E_3 = III^{rd}$ onisation Potential

- I^{st} Ionisation Potential $< II^{nd}$ Ionisation Potential
- < IIIrd onisation Potential

- (b) Electron can not be removed from solid state of an atom, it has to convert in gaseous form, Energy required for conversion from solid state to gaseous state is called Sublimation energy.
- (c) Ionisation Potential is always an endothermic process $(\Delta H = +ve)$
- (d) It is measured in eV/atom (electron volt/atom) or Kcal/mole or KJ/mole
- (iii) Factors affecting ionisation potential :
- (A) **Atomic size :** Larger the atomic size, smaller is the Ionisation Potential It is due to that the size of atom increases the outermost electrons farther away from the nucleus and nucleus loses the attraction on that electrons and hence can be easily removed.

Ionisation Potential
$$\propto \frac{1}{\text{Atomic size}}$$

(B) Effective nuclear charge (Z_{eff}) : Ionisation potential increases with the increase in nuclear charge between outermost electrons and nucleus.

Ionisation Potential 🗴 Effective nuclear charge

(C) Screening effect : Higher is the screening effect on the outer most electrons causes less attraction from the nucleus and can be easily removed, which is leading to the lower value of Ionisation Potential



- (D) Penetration power of sub shells :
- (a) Order of attraction of subshells towards nucleus (Penetration power) is -

s > p > d > f

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

Ex.

Be B $1s^2, 2s^2$ $1s^2, 2s^2 2p^1$

Ionisation Potential Be > B

After loosing one e⁻, B attains electronic configuration

of Be, so IInd ionisation potential of B is more than

Be. II^{nd} Ionisation Potential of B > Be

- (E) Stability of half filled and fully filled orbitals :
- (a) Half filled p³,d⁵, f⁷ or fully filled s², p⁶, d¹⁰, f¹⁴ are more stable than others so it requires more energy.

Ex. N

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

0

 I^{st} Ionisation Potential order is O < NBecause of half filled p-orbitals in N, its ionisation energy (stability) is higher than O.

 I^{st} ionisation potential order Na < Al < Mg

(b) Because s-orbital in Mg is completely filled and its penetration power is also higher than p-orbital (Al).
 IInd ionisation potential order

Mg^+	<	Al^{+}	<	Na^+
(2,8,1)		(2,8,2)		(2,8)

(F) Oxidation state :

Ionisation Potential \propto	oxidation	state of an	atom
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Ion with high oxidation state will have high ionisation potential. Ex. $Fe^{+3} > Fe^{+2} > Fe$

(iv) In a group : Size increase so ionisation potential decrease.

Li Na K Rb Cs

Size increases, Ionisation Potential decreases Exception :

• Ionisation Potential Ga > Al (While Ionisation Potential decreases down the group it is due to

Transition contraction)

•

Ionisation Potential Hf > Zr 5d 4d j

(While Ionisation Potential should decreases down the group. It is due to lanthanide contraction)

(v) In a period : In a period atomic size decreases and zeff increases so removal of electron becomes difficult and ionisation potential increases.

Li Be B C N O F Ne

atomic size decreases, zeff increases, Ionisation Potential increases.

(vi) Application of ionisation potential :

(A) Metallic and non metallic character :

Metallic \longrightarrow Ionisation Potential Low (Na, K, Rb etc.)

non metallic \longrightarrow Ionisation Potential High (F, Cl, Br etc.)

Ionisation Potential $\propto \frac{1}{\text{Metallic property}}$

(B) Reactivity :



Reactivity increases down the group as ionisation potential decreases.

(C) Reducing character :



- (a) IA group has minimum ionisation potential so they are strong reducing agents in gaseous state (Li < Na < K < Rb < Cs)
- (b) IA group In Aqueous state

reducing character $Li > K \ge Rb > Cs > Na$ As the degree of hydration is more in Li due to high charge density.

(D) Stability of oxidation states :

- (a) If the difference between two successive ionisation potential ≥ 16 eV then lower oxidation state is stable.
- **Ex.** Na \longrightarrow Na⁺ Ist ionisation potential Na⁺ \longrightarrow Na⁺² IInd ionisation potential } 42.7 eV Difference of ionisation potential > 16 eV So Na⁺ is more stable.
- (b) If the difference between two successive ionisation potential ≤ 11 higher oxidation state is stable.

then the hydroxide will be acidic in nature.

Ex. Mg \longrightarrow Mg⁺ Istionisation potential Mg⁺ \longrightarrow Mg⁺² IInd ionisation potential $\}$ 7.4 eV Difference of ionisation potential < 11 eV So Mg⁺² is more stable.

$$Al \longrightarrow Al^{\scriptscriptstyle +}$$

$$\begin{array}{l} Al \longrightarrow Al \\ Al^{+} \longrightarrow Al^{+2} \\ Al^{+2} \longrightarrow Al^{+3} \end{array} \right\} \begin{array}{l} 12.8 \text{ eV So } Al^{+} \text{ is more} \\ 6.0 \text{ eV So } Al^{+3} \text{ is more stable} \end{array}$$

Al⁺ is stable only in gaseous state

[{]Al⁺³ is stable in liquid and solid state.

(E) Basic nature : It is property of elements with loosely held electrones

Ionisation Potential $\propto \frac{1}{\text{Basic Property}}$

 $M_2O > MO$, $Na_2O > MgO$, $NaOH > Mg(OH)_2$ $Cs_2O > Rb_2O > K_2O > Na_2O > Li_2O$

- (F) To determine the number of valence electron :
- Ex. LiIst IonisationIInd IonisationPotentialPotential5.4 e V75 eV

High energy difference in two successive ionisation potential shows change in energy levels ie. valence shell has only one electron.

(G) Ionic nature : Ionic nature of elements decreases from left to right in a period.



(H) Nature of hydroxide :

In a hydroxide AOH – Nature of hydroxide depends on ionisation potential of A.

A–O–H. If ionisation potential of A is lower than oxygen, it will loose electron and will be basic

$$A-OH \longrightarrow A^+ + OH^-$$

If A has high ionisation potential than oxygen – $A - O - H \longrightarrow AO^{-} + H^{+}$

 $\begin{array}{c|cccc} NaOH & Mg(OH)_2 & Al(OH)_3 & Si(OH)_4 & P(OH)_3 & SO_2(OH)_2 & ClOH \\ \hline \\ \hline \\ Ionisation Potential increases & Acidic nature increases & Basic nature decreases \\ \hline \end{array}$

Note: Ionisation potential decreases in a group so, basic nature increases.

ELECTRON AFFINITY/ELECTRON GAIN ENTHALPY

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

 $X + e^{-} \longrightarrow X^{-} +$ Electron Affinity

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

 $X + e^- \longrightarrow X^- + I^{st}$ Electron Affinity

(iii) If Ist Electron affinity is exothermic

 Δ_{eg} H (electron gain enthalpy) = -ve Ist Electron affinity is mostly positive

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

 $\{Energy evolved in Electron affinity to convert neutral atom to anion\} = \{Energy absorbed to remove e⁻ from anion\}$

(vii) Factors affecting electron affinity :

- **Ex.** X + $e^- \xrightarrow{\text{Electron Affinity}} X^- + \text{Energy}$
- (v) On adding 2nd electron in anion X⁻ $X^- + e^- \longrightarrow X^{-2} - II^{nd}$ Electron Affinity or X⁻ + e⁻ + II^{nd} Electron Affinity $\longrightarrow X^{-2}$ Δ_{eg} H (2nd electron gain enthalpy) = + ve
- (vi) IInd Electron affinity is endothermic process, i.e. energy has to given to introduce 2nd electron in an anion, because of repulsion between negative charge on anion (X⁻) and 2nd electron.IInd Electron Affinity is always negative.
- Ex. $O(gas) + e^- \longrightarrow O^-(gas) + 168$ K.cal (exothermic) $O^- + e^- \longrightarrow O^{-2} - 912.7$ K.cal

Net reaction O(g) + 2e⁻ \longrightarrow O⁻²(g) Electron Affinity = -744.7 K.cal (endothermic)

(A)	Atomic size	Electron Affinity $\propto \frac{1}{\text{Atomic size}}$
(B)	Screening effect	Electron Affinity $\propto \frac{1}{\text{Screening effect}}$
(C)	Effective nuclear charge (Z_{eff})	Electron Affinity ∝ Zeff

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

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

Ex.

- (i) Elements (He, Ne, Ar , ..), Electron Affinity= Zero
- (ii) Elements (Be, Mg, Ca,), Electron Affinity=Approximately Zero
- (iii) Elements (N, P, As,), ElectronAffinity= Very less

(viii) In period:

- (a) Electron affinity increases along the period due to increase in zeff and decrease in atomic size.
- (b) Halogens possess maximum electron affinity due to small size and maximum zeff and after gaining one electron. They attain a stable inert gas configuration.

In 2^{nd} period – Ne < Be < N < B < C < O < F In 3^{rd} period – Ar < Mg < P < Al < Si < S < Cl

(ix) In Group :

- (a) Electron affinity decreases down the group. As the size increase and attraction on additional electron is less.
- (b) Electron affinity of 3rd period element is greater than electron affinity of 2nd period elements of the respective group.

F Cl 1s², 2s²2p⁵ [Ne] 3s²3p⁵

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
S > O > Se	Si > C > P > N

Note :

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

(x) Application of Electron affinity :

(a) Electron affinity \propto Oxidising nature

But F has more oxidising power than Cl because F.

(i) has high standard reduction potential.

(ii) Has high hydration enthalpy.

(iii)Low bond dissociation energy.

(xi) Difference between electronegavity and Electron Affinity :

Electronegativity	Electron Affinity
Tendency of an atom in a molecule	Energy released when an electron is added to
to attract the bonded electrons	neutral isolated gaseous atom
It regularly changed in a period	It does not changes regularly
It has no unit	It is measured in eV/ atom or KJ mol ⁻¹ or K cal mole ⁻¹

ELECTRONEGATIVITY (EN)

- (i) The tendency of an atom to attract shared electrons towards itself is called electronegativity.
- Electronegativity and Electron affinity both have tendency to attract electrons but electron affinity is for isolated atoms. Where as electronegativity is for bonded atoms.
- (iii) A polar covalent or ionic bond of A B may be broken as

(a) $A-B \longrightarrow A^-$: + B^+ (Electronegativity A > Electronegativity B)

or (b) $A - B \longrightarrow A^+ + :B^-$

(ElectronegativityA<ElectronegativityB) depending on their tendency to attract bonded electron.

- (iv) There is no unit of electronegativity as electronegativity is tendency of a bonded atom not an energy
- (v) Pauling explained it first time.
- (vi) Electronegativity of some other elements are as follows –

						H 2.1	In Pauling's scale, elements having almost same electronegativity are-
Li	Be	В	С	N	0	F	C = S = I = 2.5
1.0	1.5	2.0	2.5	3.0	3.5	4.0	N - Cl - 3.0
Na	Mq	Al	Si	Р	S	Cl	N = CI = 0.0
0.9	1.2	1.5	1.8	2.1	2.5	3.0	P = H = 2.1
Κ						Br	$C_S = Fr = 0.7$
0.8						2.8	Be = AI = 1.5
Rb						Ι	
0.8						2.5	
Cs							
0.7							
Fr							
0.7							

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

(vii) Factors Affecting electronegativity :

(A)	Atomic size	(B)	Effective nuclear charge ($\mathbf{Z}_{_{\mathrm{eff}}}$)
	Electronegativity $\propto \frac{1}{\text{Atomic size}}$		Electronegativity ∝ Zeff
(C)	Hybridisation state of an atom	(D)	Oxidation state
	Electronegativity \propto % s character in hybridised atom	l	Electronegativity \propto oxidation state
	$sp > sp^2 > sp^3$		$Mn^{+2} < Mn^{+4} < Mn^{+7}$
	s character 50% 33% 25%		O^{-2} < O^{-1} < O < O^{+1} < O^{+2}
	Electronegativity 3.25 2.75 2.5		$Fe < Fe^{+2} < Fe^{+3}$

Key Points :

- s-orbital is nearer to nucleus so by increasing scharacter in hybridisation state, Electronegativity also increases.
- Electronegativity does not depends on filled or half filled orbitals, because it is a tendency to attract bonded electron, not to gain electron from out side.

(viii) Periodic table & Electronegativity :

- (a) Electronegativity decreases down the group.
- (b) In period on moving from left to right electronegativity increases.

Exceptions : In '0' group – Electronegativity is always zero, because inert gas do not form molecule.

- (c) Electronegativity of Cs and Fr are equal, it is because from ${}_{55}$ Cs to ${}_{87}$ Fr only one shell increases but nuclear charge (No. of proton) increases by +32.
- (d) So effect of nuclear charge balanced the effect of increase in number of shell.

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

So Fluorine is called **Black sheep** element.

- (e) In group of IIB elements (Zn, Cd, Hg) value of electronegativity increases down the group, because of lanthanide contraction
- (f) In IIIA group, value of electronegativity increases down the group, because of transition contraction Electronegativity of Ga > Electronegativity of Al

(ix) Application of electronegativity :

(A) Metallic and non metallic nature :

Low electronegativity \longrightarrow Metals

High electronegativity → Non metals

Metallic character increases down the group but decreases along a period.

(B) **Bond length :**

 $\Delta EN \propto \frac{1}{Bond length}$ Here $\Delta EN = difference in$

electronegativities of bonded atoms

Ex. HF < HCl < HBr < HI

(D) Acidic strength of hydrides :

Bo	ond energy (Strength)	x	stability of mo	lecule
Ord	er of stability of hyd	lroh	alides	and
HF >	> HCl > HBr > HI			
In V	A group –			
NH ₃	3			PH_3

Thermal stability decreases

(E) Reactivity :

Bond energy \propto Stability $\propto \frac{1}{\text{Reactivity}}$
Bond energy 🗴 Difference in electronegativities
Difference in electronegativities \propto Stability $\propto \frac{1}{\text{Reactivity}}$

HF HCl HBr

 $\Delta\, EN$ decreases, stability decreases, Reactivity increases

HI

Note :

HI is most reactive hydrohalides or strongest acid among all hydrohalides.

(F) Nature of bonds :

(a) According to Hanny & Smith formula

Ionic % = $16 (X_A - X_B) + 3.5 (X_A - X_B)^2$ Here X_A = Electronegativity of A X_B = Electronegativity of B HF has minimum bond length because of much difference in the electronegativities of H and F.

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

Bond energy \propto Electronegativity difference

HF > HCl > HBr > HI

order of acidic strength HF < HCl < HBr < HI

AsH₃

Acidic character increases

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

(b) According to Gallis

$$\begin{split} X_{A} - X_{B} &\geq 1.7 & \text{Ionic} \\ X_{A} - X_{B} &\leq 1.7 & \text{Covalent} \\ \text{If } X_{A} &= X_{B} \text{; then } A - B \text{ will be non polar.} \end{split}$$

- Ex. H—H, F—F If $X_A > X_B$ and difference of electronegativities is small then
- **Ex.** H_2O (H^{δ_+} ---- O^{δ_-} ---- H^{δ_+}) If $X_A >> X_B$ and $X_A - X_B$ difference of electronegativities is high then

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

Ex. Na⁺Cl⁻

In HF, $X_A - X_B = 1.9$, which is more than 1.7, even then it is covalent compound.

(G) Nature of hydroxides :

- (a) As per Gallis, In AOH if electronegativity of A is more than 1.7 (Non metal) then it is acidic in nature.
- (b) If electronegativity of 'A' is less than 1.7 (metal) then AOH will be basic in nature

NaOH

0.9

Ex.

Nature

Basic (c) If $X_{A} - X_{0} \ge X_{0} - X_{H}$

3.0 Acidic then AO bond will be more polar and will break up as

ClOH

- A OH \longrightarrow A⁺ + OH⁻ It shows basic nature
- (d) If $X_A X_O \leq X_O X_H$ $A - O \longrightarrow H^+ + AO^-$ It shows Acidic nature

Ex. In NaOH

 $X_0 - X_{Na}(2.6) > X_0 - X_H(1.4)$ So hydroxide is basic

In ClOH

 $X_0 - X_{C1}(0.5) < X_0 - X_H(1.4)$ So hydroxide is acidic

(H) Nature of oxides : Consider an oxide AO

 $X_{A} - X_{O} > 2.3$ Basic oxide If $X_{A} - X_{O} = 2.3$ Amphoteric oxide If $X_A - X_O < 2.3$ Acidic oxide If

(a) Along a period acidic nature increases.

(b) Down the group basic nature increases C N O F Li Be В Si P S Cl Na Mg Al

 $X_A - X_O > 2.3$ $X_A - X_O = 2$ Basic

$$X_{\rm A} - X_{\rm O} < 2.3$$

Amphoteric Acidic

ie. when in periodic table the distance between the element and oxygen increases, basic character increases.

 $NO_2 > ZnO > K_2O$

acidic character decreases

Note:

BeO, Al₂O₃, ZnO, SnO, PbO, SnO₂, PbO₂, Sb₂O₃ etc. are amphoteric oxides.

CO, H₂O, NO, N₂O etc. are neutral oxides.

Acidic strength of	oxide and oxyacid	∞	Electronegativity
B_2O_3	CO ₂		N ₂ O ₅

EN increase, acidic nature increasse.

 $SO_3 > SeO_3 > TeO_3$ $H_2SO_4 > H_2SeO_4 > H_2TeO_4$ $H_2SO_3 > H_2SeO_3 > H_2TeO_3$ $HNO_3 > H_2CO_3 > H_3BO_3$ $H_3PO_4 > H_3AsO_4 > H_3SbO_4$ HOF > HOCl > HOBr > HOI $HClO_4 > HBrO_4 > HIO_4$ $N_2O_5 > P_2O_5 > As_2O_5$ $N_2O_3 > P_2O_3 > As_2O_3 > Sb_2O_3$

Acidic nature oxidation state

Acidic properties increases with increasing oxidation state of an element

$$\begin{aligned} &\text{HClO}_4 > \text{HClO}_3 > \text{HClO}_2 > \text{HClO}\\ &\text{H}_2\text{SO}_4 > \text{H}_2\text{SO}_3\\ &\text{SO}_3 > \text{SO}_2 \end{aligned}$$

$$HNO_{3} > HNO_{2}$$
$$N_{2}O_{5} > N_{2}O_{3}$$
$$Sb_{2}O_{5} > Sb_{2}O_{2}$$

(I) Hydrolysis of AX :

Where A = O there element and X = Halogen

(a) If electronegativity of X > Electronegativity of A then on hydrolysis product will be HX.

Ex. (BCl₂), Electronegativity of Cl>Electronegativity of B

$$B^{\delta^{+}} \underbrace{Cl^{\delta^{-}}}_{Cl^{\delta^{-}}} + H^{+} \underbrace{OH^{-}}_{H^{+}} \longrightarrow 3HCl + B(OH)_{3}$$

$$PCl_{3} + 3HOH \longrightarrow 3HCl + H_{3}PO_{3}$$

$$PCl_{3} \cdot Cl_{2} + H_{2}O \longrightarrow 2HCl + POCl_{3}$$

$$POCl_{3} + 3HOH \longrightarrow 3HCl + H_{3}PO_{4}$$

- (b) If electronegativity of $X \le electronegativity of A$ then on hydrolysis product will be HOX (hypohalous acid)
- **Ex.** Cl_2O $O^{-} < Cl^{\delta^+}$ Here electronegativity of O > Cl

So On hydrolysis –

$$O^{\delta^{-}} \left\{ \begin{array}{ccc} Cl^{\delta^{+}} & HO^{-} & H^{+} \\ Cl^{\delta^{+}} & HO^{-} & H^{+} \end{array} \right\} \longrightarrow 2HOCl + H_{2}O$$

(J) Nomenclature of inorganic compounds :

Prefix ——less electronegative element

Suffix —— More electronegative element

OCl,(Wrong) **Ex.** Cl₂O(Right)

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

OF, Oxygen difluoride

- ICl Iodine chloride
- Exception (Here H is less electronegative NH, but it is suffix)
- (x) Electronegativity Scale :
- (A) Pauling Scale (Bond energy scale) :

 $0.208\sqrt{\Delta AB} = X_A - X_B$

Here \triangle AB is resonance energy of AB molecule This equation gives difference in electronegativity values.

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

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

Relation between Pauling scale & Mulliken scale

 $X_{p} = 0.336 (X_{m} - 0.615)$

- $X_{p} =$ Electronegativity given by pauling $X_{m} =$ Electronegativity given by mulliken
- ²^[2] If ionisation potential and electron affinity are given in eV, then electronegativity by Mulliken on Pauling scale will be

v	_	Ionisation Potential + Electron Affinity
Λ _m	_	5.6

2/21 If ionisation potential and electron affinity are given in K.cal/mole then

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

(C) Allred Roschov's scale: As per Roschov, electronegativity is the force by which nucleus of an atom attracts electron which are on the covalent radius.

$$(X_{AR}) \text{ Electronegativity} = \frac{Z_{eff} \times e^2}{r^2} \text{ or}$$
$$X_P = 0.359X_{AR} + 0.744$$