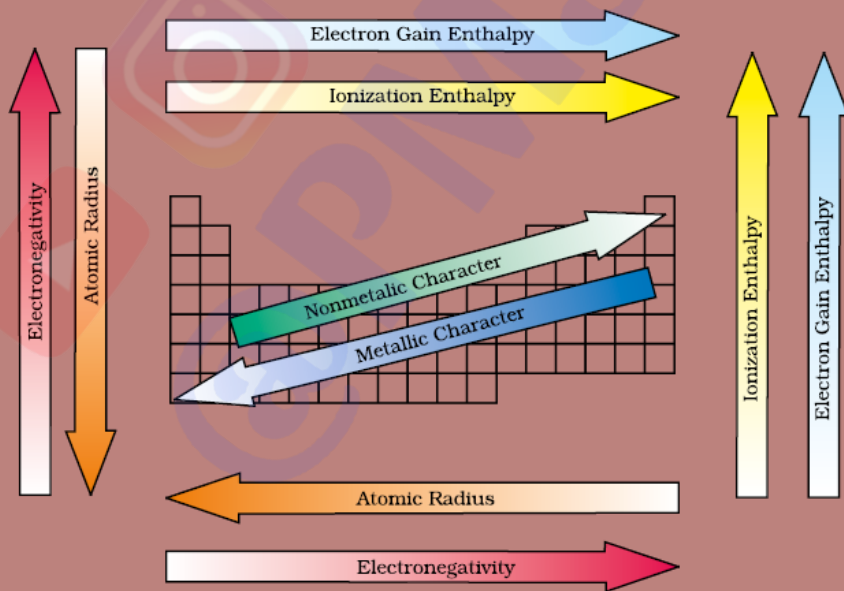


PERIODIC PROPERTIES



Q:- Why do we need to classify elements?

<u>Ans:-</u>	Year	→	no. of elements discovered
	1800		31
	1865		63
	At present		118

It is very difficult to study all the properties of such large number of elements individually. So scientists searched for a systematic way to organise these elements.

Q:- What is periodic Table?

Ans:- Periodic Table is an systematic arrangement of all the elements known to man in accordance with their increasing atomic number and recurring chemical and physical properties. They are arranged in tabular formate to keep similar property elements in vertical column (known as groups). Horizontal rows are known as periods.

* Genesis of Periodic Classification:-

(i) Dobereiner's Law of Triads:

- He formed group of 3 elements in such a way that, when the elements are arranged according to increasing atomic mass, then atomic mass of middle element is average of atomic masses of first and last element.
- He found that elements of such groups have similar properties. Such groups were known as Dobereiner's Triad.
- **Mathematically:** If elements A, B, C have atomic masses x, y, z respectively, then $(y = \frac{x+z}{2})$, also A, B, C have similar properties.

eg: Li (A=7), Na (A=23), K (A=39).

$$\text{atomic mass of Na} = 23 = \frac{7+39}{2} = \text{average}$$

More examples of Dobereiner's Triad:-

Ca (A=40), Sr (A=88), Ba (A=137).

U (A=35.5), Br (A=80), I (A=127).

P (A=31), As (A=75), Sb (A=120).

S (A=32), Se (A=79), Te (A=127).

Trick:- If atomic numbers of elements A, B and C are x, y and z , and $(y-x) = (z-y)$, then A, B and C are in triad. (no need of atomic mass).



eg → Li (Z=3), Na (Z=11), K (Z=19) ⇒ $11-3 = 19-11 = 8$
Dobereiner's Triad.

eg → Ca (Z=20), Sr (Z=38), Ba (Z=56) ⇒ $38-20 = 56-38 = 18$
Dobereiner's Triad.

Drawback:

- He wasn't able to classify all the known elements into triads.
- Properties of elements of different triads can't be compared with each other.
- He made some triads based on similarities of properties but do not follow average of atomic mass rule, such as:-

(i) H ($Z=1$), Li ($Z=7$), Na ($Z=23$)

$$\text{average} = \frac{23+1}{2} = 12 \neq \text{atomic mass of Li}$$

(ii) O ($Z=16$), S ($Z=32$), Se ($Z=79$)

$$\text{average} = \frac{16+79}{2} = 47.5 \neq \text{atomic mass of S}$$

(iv) A.E.B. de Chancourtois :

- He attempted to classify elements in order of increasing atomic weight and made a cylindrical table, which was used to display periodic recurrence of properties.
- Drawback: He only published his paper but did not publish his actual graph with irregular arrangements

(v) Newland's Law of Octaves:-

- He arranged the elements in increasing order of their atomic weights and observed that every eighth element had properties similar to the first element. Just like every eighth note that resembles the first note in octaves of music.

Do	Re	Mi	Fa	So	La	Ti
सा	रे	मि	फा	सो	ला	ति
Li	Be	B	C	N	O	F
Na	Mg	Al	Si	P	S	Cl
K	Ca	---	---	---	---	---
---	---	---	---	---	---	---

Drawback:

- His Law seemed to be true only for elements upto to calcium.
- Untill then noble gases weren't discovered, after discovery of noble gases, his law was modified. That is, properties of every ninth element are similar to properties of first one.

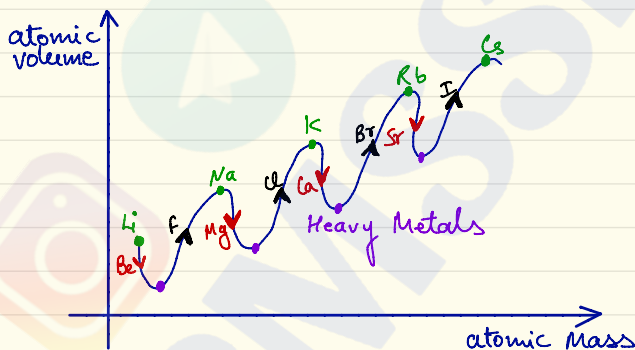
(vi) Lothar Meyer's Curve:

- He plotted the physical properties such as atomic volume, melting point and boiling point against atomic weight and obtained a periodically repeated pattern.

→ He observed that elements having similar properties occupy similar positions.

→ In the following graph of atomic volume vs atomic mass, he observed that :

- i) Alkali metals occupy maxima of curve.
- ii) Alkaline earth metals occupy midpoints of descending curve.
- iii) Halogens occupy midpoints of ascending curve.
- iv) Heavy Metals occupy minima of curve.



Drawbacks:

- Graphical method was hard to study.
- He also developed a table but before he got it published, Mendeleev's Periodic Table was already published.
- It was difficult to find similar positions on a curve of different properties as periodicity and nature of graph was not similar.

(vii) Mendeleev's Periodic Table:

- He was responsible for publishing the Periodic Law for the first time. It states as follows:-
Atomic mass is the fundamental property of elements and all other physical and chemical properties are periodic function of atomic mass.
- He fully recognised the significance of periodicity and used broader range of physical and chemical prop. to classify elements.
- His main basis was the similarities in empirical formulas and properties of compounds formed by the elements.

Characteristics:

- He arranged 63 known elements into:
 - 12 horizontal rows \Rightarrow Series
 - 8 vertical columns \Rightarrow Groups.
- Groups were further divided:
 - Subgroup A: Normal Elements.
 - Subgroup B: Transition Elements.
- Some of 12 series were grouped horizontally, these horizontal groups were known as periods. (7 periods)
- The VIII group consist of 9 elements in three vertical columns.

Merits:

- He classified all known elements into groups according to their similar properties. Hence, study of elements was made easier due to tabular classification.
- He realized that some elements did not fit in right place according to other properties, if placed according to atomic weight. So he ignored atomic weights, thinking that measurements might be incorrect.

Hence, led to corrections of atomic weights of some elements - U, Be, In, Au, Pt.

- As he arranged elements with similar property, he proposed that some elements were still undiscovered, and left several gaps, which helped in discovery of new elements.

Eka Aluminium (Al) \rightarrow Ga (Aaloo - Galoo)

Eka Silicon (Si) \rightarrow Ge (Sita - Geeta)

Eka Boron (B) \rightarrow Sc (B. Sc)

Eka Manganese (Mn) \rightarrow Tc (M. Tech)

Demerits:

- Position of Hydrogen was not fixed.
- Position of isotopes was not clear.
- He could not explain cause of periodicity.
- Some pairs of elements do not follow the increasing order of atomic weights, k/a Anomalous Pairs.

$\left(\begin{array}{cc} \text{Te} & \text{I} \\ 127.5 & 127 \end{array} \right)$	$\left(\begin{array}{cc} \text{Co} & \text{Ni} \\ 58.9 & 58.6 \end{array} \right)$	$\left(\begin{array}{cc} \text{Th} & \text{Pa} \\ 232 & 231 \end{array} \right)$
--	---	---

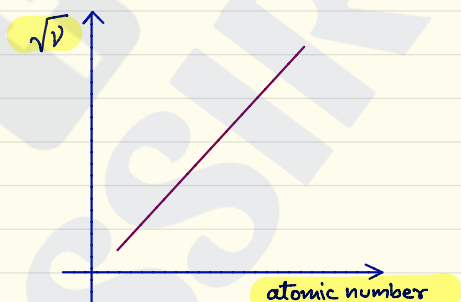
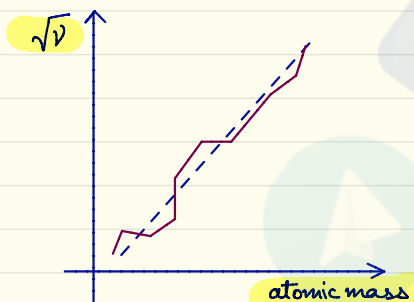
जेटी इच्छा Coke जेटी ठंडा Pepsi

SERIES	GROUPS OF ELEMENTS								
	0	I	II	III	IV	V	VI	VII	VIII
1	-	Hydrogen H 1.008	-	-	-	-	-	-	-
2	Helium He 4.0	Lithium Li 7.03	Beryllium Be 9.1	Boron B 11.0	Carbon C 12.0	Nitrogen N 14.04	Oxygen O 16.00	Fluorine F 19.0	-
3	Neon Ne 19.9	Sodium Na 23.5	Magnesium Mg 24.3	Aluminium Al 27.0	Silicon Si 28.4	Phosphorus P 31.0	Sulphur S 32.06	Chlorine Cl 35.45	-
4	Argon Ar 38	Potassium K 39.1	Calcium Ca 40.1	Scandium Sc 44.1	Titanium Ti 48.1	Vanadium V 51.4	Chromium Cr 52.1	Manganese Mn 55.0	Iron Fe 55.9
5	-	Copper Cu 63.6	Zinc Zn 65.4	Gallium Ga 70.0	Germanium Ge 72.3	Arsenic As 75	Selenium Se 79	Bromine Br 79.95	Cobalt Co 59
6	Krypton Kr 81.8	Rubidium Rb 85.4	Strontium Sr 87.6	Yttrium Y 89.0	Zirconium Zr 90.6	Niobium Nb 94.0	Molybdenum Mo 96.0	-	Ruthenium Ru 101.7
7	-	Silver Ag 107.9	Cadmium Cd 112.4	Indium In 114.0	Tin Sn 119.0	Antimony Sb 120.0	Tellurium Te 127.6	Iodine I 126.9	Rhodium Rh 103.0
8	Xenon Xe 128	Caesium Cs 132.9	Barium Ba 137.4	Lanthanum La 139	Cerium Ce 140	-	-	-	Palladium Pd 106.5
9	-	-	-	-	-	-	-	-	-
10	-	-	-	Ytterbium Yb 173	-	Tantalum Ta 183	Tungsten W 184	-	Osmium Os 191
11	-	Gold Au 197.2	Mercury Hg 200.0	Thallium Tl 204.1	Lead Pb 206.9	Bismuth Bi 208	-	-	Iridium Ir 193
12	-	-	Radium Ra 224	-	Thorium Th 232	-	Uranium U 239	-	Platinum Pt 194.9
	R	R ₂ O	RO	R ₂ O ₃	RO ₂ RH ₄	HIGHER SALINE OXIDES R ₂ O ₅ RO ₃ R ₂ O ₇ HIGHER GASEOUS HYDROGEN COMPOUNDS RH ₃ RH ₅ RH ₇			RO ₄

(vii) Modern Periodic Law & The Modern Periodic Table:

Henry Mosley's Experiment:

→ Mosley Measured frequency of x -rays emitted by metal, when it is bombarded with high speed e^- . and plotted these 2 graphs:



→ He observed that the curve of $\sqrt{\nu}$ vs atomic mass is approximately a straight line but $\sqrt{\nu}$ vs at. number is perfectly a straight line.

→ So, He concluded that instead of atomic mass, atomic number should be fundamental property.

→ Neil's Bohr, considered Mosley's suggestion and modified Mendeleev's periodic law, to give:

Modern Periodic Law:

Atomic Number is the fundamental property of elements and all other physical and chemical properties are periodic functions of atomic no.

* Modern Periodic Table / Long form of Periodic Table:

→ Modern periodic table is constructed on the basis of modern periodic law and electronic configuration.

→ It consists of : 18 vertical columns \Rightarrow Groups
7 horizontal rows \Rightarrow Periods
no. of elements \Rightarrow 118.

→ Orbitals filled in " n^{th} " period :
(ns) ... (n-3)g, (n-2)f, (n-1)d, np.

→ No. of elements in " n^{th} " period:

Period	orbitals filled	no. of elements	type of period
1	1s	$1 \times 2 = 2$	Shortest
2	2s, 2p	$4 \times 2 = 8$	Short
3	3s, 3p	$4 \times 2 = 8$	Short
4	4s, 3d, 4p	$9 \times 2 = 18$	long
5	5s, 4d, 5p	$9 \times 2 = 18$	long
6	6s, 4f, 5d, 6p	$16 \times 2 = 32$	longest
7	7s, 5f, 6d, 7p	$16 \times 2 = 32$	longest

↑
Magic No.s

→ Mathematically :

no. of elements in n^{th} period:

$$\text{if } n \text{ is even} = 2 \left(\frac{n+2}{2} \right)^2$$

$$\text{if } n \text{ is odd} = 2 \left(\frac{n+1}{2} \right)^2$$



Trick



grp1: He Li Na K Rb Cs Fr
He Lina ki Rab Se Jariyad.

grp2: Be Mg Ca Sr Ba Ra
Beta Mange Car Scooter Bapu Raji.

grp3: Sc Y La Ac
Scooter Yamaha Lao Esi.

grp4: Ti Zr Hf
Tea jara half.

grp5: V Nb Ta
Vaah Nobita

grp6: Cr Mo W
Corona Maut Wala

grp7: Mn Tc Re
Man Tak Rahe

grp8: Fe Ru Os
Ferrous

grp9: Co Rh Ir
Courier

grp10: Ni Pd Pt
Nahi Padode Pitoge

grp11: Cu Ag Au
Kyu Aage Aau

grp12: Zn Cd Hg
Zindagi Cadbury Hogi

grp13: B Al Ga In Tl
Bengain Aaloo Gajar in Thela

grp14: C Si Ge Sn Pb
Chemistry Sir Gines Sauki Problem.

grp15: N P As Sb Bi
Nana latekar Asmani Sab Bimar

grp16: O S Se Te Po
Oye Sanam Se Tera bol Khol dunga.

grp17: F Cl Br I At
full colour Brightness in Akai TV.

grp18: He Ne Ar Kr Xe Rn
Henna Neelam Aur Kareena ka X-ray Rangeen

3d-series: Sc Ti V Cr Mn Fe Co Ni Cu Zn
Science Teacher Vinita Kumari Mann Fe Co Ni Kyu jaan
jalati ho.

4d-series: Y Zr Nb Mo Tc Ru Rh Pd Ag Cd
Yaari Zara Nibha Maut Tak Rukavate Rah main Pade to Aag hein Cud
Ja.

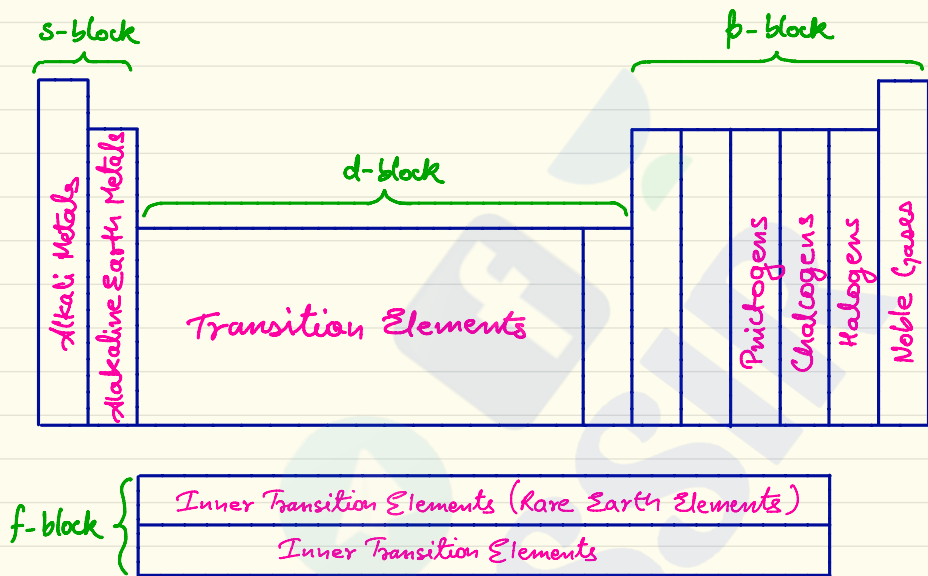
5d-series: La Hf Ta W Os Ir Pt Au Hg.
Laa Hafta Verna Osama Se Idhar Pitai Aur Hogi

6d-series: Ac Rf Db Sg Bh Hs Mt Ds Rg Cn.
Ac Main Betha Rutherford dub gaya sagar main, Bhai
Has Mat, dosa Kha raha Rahul Gandhi Kone Main.

4f-series: Ce Pr Nd Pm Sm Eu Gd Tb Dy Ho Er Tm Yb Lu.
Sir par Nadiyan prem ki samayi ya gadgad tab
dil hua arre Tumhe To Yubhi loot lu.

5f-series: Th Pa U Np Pu Am Cm Bk Cf Es Fm Md No Lr.
Tumhe padau newspaper purana, Aam Kam Bike,
Cafe main aish farmate Madona - Lara.

* Modern Periodic Table can be classified into four blocks:



① S-block:- Elements in which last e^- enters in s-subshell of valence shell, are placed in s-block.

	1	2
	IA	IIA
	3	4
	Li	Be
	$2s^1$	$2s^2$
	11	12
	Na	Mg
	$3s^1$	$3s^2$
	19	20
	K	Ca
	$4s^1$	$4s^2$
	37	38
	Rb	Sr
	$5s^1$	$5s^2$
	55	56
	Cs	Ba
	$6s^1$	$6s^2$
	87	88
	Fr	Ra
	$7s^1$	$7s^2$

Alkali Metals

Alkaline Earth Metals

- General E.C. → ns^{1-2}
- General oxidation state of grp-1 and grp-2 is +1 and +2 respectively.
- Larger in size.
- Soft metals and silvery white in colour
- Mostly forms ionic comp.
- Have low IP and highly reactive. so, never found in pure form.
- Strong Reducing Agents
- Gives flame test (except Be and Mg).

② p-block:- Elements in which last e^- enters in p-subshell of valence shell, are placed in p-block.

General E.C. $\rightarrow ns^2 np^{1-6}$

$ns^2 np^1$	$ns^2 np^2$	$ns^2 np^3$	$ns^2 np^4$	$ns^2 np^5$	$ns^2 np^6$
13	14	15	16	17	18
B	C	N	O	F	Ne
Al	Si	P	S	Cl	Ar
Ga	Ge	As	Se	Br	Kr
In	Sn	Sb	Te	I	Xe
Tl	Pb	Bi	Po	At	Rn
Nh	Fl	Mc	Lv	Ts	Og

\rightarrow Comprise of group 13 to 18.

\rightarrow s-block elements and p-block elements together known as Representative elements or Main Group Elements.

\rightarrow Good oxidising agents.

\rightarrow Mostly forms covalent bonds.

\rightarrow Harder and heavier than s-block elements and have high E.A.

\rightarrow p-block elements are mostly non-metals and metalloids but few are Metals. (Metalloids \rightarrow Ge, As, Se, Te, Sb)

\rightarrow H, N, O, F, Cl and noble gases exist in gas form at room temp.

\rightarrow Br is the only non-metal which exist in liquid form.

\rightarrow Grp 15 - Pnictogens: refers to choking / suffocating property.

Grp 16 - Chalcogens: ore forming elements.

Grp 17 - Halogens: sea-salt forming elements.

Grp 18 - Noble / Inert gas: Chemically inert elements.

③ d-block:- Elements in which last e^- enters in d-subshell of penultimate shell, are placed in d-block.

	3	4	5	6	7	8	9	10	11	12
	III A	IV A	V A	VIA	VII A	← VIII →			I B	II B
I st Transition Series 3d - series	21 Sc $3d^4 4s^2$	22 Ti $3d^2 4s^2$	23 V $3d^3 4s^2$	24 Cr [*] $3d^5 4s^1$	25 Mn $3d^5 4s^2$	26 Fe $3d^6 4s^2$	27 Co $3d^7 4s^2$	28 Ni $3d^8 4s^2$	29 [*] Cu [*] $3d^{10} 4s^1$	30 Zn $3d^{10} 4s^2$
II nd Transition Series 4d - series	39 Y $4d^1 5s^2$	40 Zr $4d^2 5s^2$	41 [*] Nb [*] $4d^4 5s^1$	42 [*] Mo [*] $4d^5 5s^1$	43 Tc $4d^5 5s^2$	44 [*] Ru [*] $4d^6 5s^1$	45 [*] Rh [*] $4d^7 5s^1$	46 [*] Pd [*] $4d^{10}$	47 [*] Ag [*] $4d^{10} 5s^1$	48 Cd $4d^{10} 5s^2$
III rd Transition Series 5d - series	57 [*] La [*] $5d^1 6s^2$	72 Hf $4f^{14} 5d^2 6s^2$	73 Ta $5d^3 6s^2$	74 W $5d^4 6s^2$	75 Re $5d^5 6s^2$	76 Os $5d^6 6s^2$	77 Ir $5d^7 6s^2$	78 [*] Pt [*] $5d^9 6s^1$	79 [*] Au [*] $5d^{10} 6s^1$	80 Hg $5d^{10} 6s^2$
IV th Transition Series 6d - series	89 [*] Ac ^{**} $6d^1 7s^2$	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 [*] Rg [*]	112 Cn

^{*} Exceptions of (n+1) rule

→ General $\Sigma.C \rightarrow ns^2 (n-1)d^{0-10}$

not Transition Elements

→ All d-block elements except group-12 elements are known as transition elements.

→ Transition Elements: Elements with incomplete d-subshell of penultimate shell, either in ground state or stable oxidation state.

→ All d-block elements are hard metals, hence show all properties of metals - good conductor of heat and electricity, malleable, ductile, sonorous etc.

→ They have variable oxidation state due to less energy gap between (n-1)d and (ns) subshells.

→ They mostly form coloured ions and show paramagnetism.

→ Coinage Metals → Cu, Ag, Au

→ Noble Metals → Au, Pt.

(4) f-block:- Elements in which last e^- enters in f-subshell of anti-penultimate shell, are placed in f-block.

Lanthanoids $4f^{0-14}5d^{0-1}6s^2$	58 Ce★	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64★ Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
	$4f^05d^06s^2$	$4f^15d^06s^2$	$4f^25d^06s^2$	$4f^35d^06s^2$	$4f^45d^06s^2$	$4f^55d^06s^2$	$4f^75d^06s^2$	$4f^85d^06s^2$	$4f^95d^06s^2$	$4f^{10}5d^06s^2$	$4f^{11}5d^06s^2$	$4f^{12}5d^06s^2$	$4f^{13}5d^06s^2$	$4f^{14}5d^06s^2$
** Actinoids $5f^{0-14}6d^{0-1}7s^2$	90★ Th★	91★ Pa★	92★ U★	93★ Np★	94★ Pu	95★ Am	96★ Cm	97★ Bk	98★ Cf	99★ Es	100★ Fm	101★ Md	102★ No	103★ Lr
	$5f^06d^07s^2$	$5f^16d^07s^2$	$5f^26d^07s^2$	$5f^36d^07s^2$	$5f^46d^07s^2$	$5f^56d^07s^2$	$5f^76d^07s^2$	$5f^86d^07s^2$	$5f^96d^07s^2$	$5f^{10}6d^07s^2$	$5f^{11}6d^07s^2$	$5f^{12}6d^07s^2$	$5f^{13}6d^07s^2$	$5f^{14}6d^07s^2$

★ Exceptions of (n+l) rule

- All f-block elements belongs to group-3, hence group-3 is longest group with 32 elements.
- General Z.C. → $ns^2(n-1)d^{0-1}(n-2)f^{1-14}$
- These are also known as Inner Transition Elements, as they are positioned between transition elements.
- 'Pm' among lanthanoids and all Actinoids are radioactive.
- Elements having atomic no. (Z) greater than U (Z=92), are man made and termed as Transuranic Elements.
- Lanthanoids are also known as Rare Earth Elements.
- General stable oxidation state of f-block elements is +3, but lanthanides also show +2 & +4, while Actinides show +3, +4, +5, +6, +7.

* Metals, Non-Metals and Metalloids:-

Metals

- Metals are hard and are usually solid at room temp. (Hg is liquid at 25°C)
- They are malleable and ductile
- They are sonorous.
- Good conductors of heat and electricity.
- They have high melting and boiling point.
- High tensile strength.

Non - Metals

- They are usually soft and gas or solid at room temp. (Br is liquid at 25°C).
- They are brittle.
- They are non-sonorous.
- They are bad conductors. (except: graphite)
- They have low melting and boiling points.
- Low tensile strength.

→ Metallic character in periodic table :

In period : Left to right decrease.

In group : top to bottom increase.

- Metalloids : Elements that show properties of both metals as well as non-metals are metalloids. or Semi - Metals.
positioned between metals and non-metals.
(Ge, As, Se, Te, Sb)

* IUPAC Naming of Elements with atomic no. > 100:

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

Atomic No.	IUPAC Name	Symbol	Elemental Name	Symbol
101	Unnilunium	Unu	Mendelevium	Md
102	Unnilbium	Unb	Nobelium	No
103	Unniltrium	Unt	Lawrencium	Lr
104	Unnilquadium	Unq	Rutherfordium	Rf
105	Unnilpentium	Unp	Dubnium	Db
106	Unnilhexium	Unh	Seaborgium	Sg
107	Unnilseptium	Uns	Bohrium	Bh
108	Unniloctium	Uno	Hassium	Hs
109	Unnilennium	Une	Meitnerium	Mt
110	Ununilium	Unn	Darmstadtium	Ds
111	Unununium	Uun	Röntgenium	Rg
112	Ununbium	Uub	Copernicium	Cn
113	Ununtrium	Uut	Nihonium	Nh
114	Ununquadium	Uuq	Flerovium	Fl
115	Ununpentium	Uup	Moscovium	Mc
116	Ununhexium	Uuh	Livermorium	Lv
117	Ununseptium	Uus	Tennesine	Ts
118	Ununoctium	Uuo	Oganesson	Og

* Group, period, Block identification:

a) If electronic configuration is given:

→ Block: given by subshell in which last e^- enters.

→ Period: largest principle quantum no.

→ Group: If it belongs to ' n^{th} ' period and

s-block, then group no. = no. of ns e^- .

p-block, then group no. = $10 + \text{no. of } (ns + np) e^-$.

d-block, then group no. = no. of $(ns + (n-1)d) e^-$.

f-block, then group no. = 3.

eg: Br ($Z=35$): $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^5$

Block = p-block (\because last e^- enters p-subshell)

Period = 4 (\because largest quantum no. is 4)

group = $10 + \text{no. of } (4s + 4p) e^-$

$$= 10 + (2 + 5)$$

$$= 17$$

b) If atomic no. (Z) is given:

→ Block: Atomic no. of noble gas (n) = 2, 10, 18, 36, 54, 86, 118.

s-block: $Z = (n+1), (n+2)$

p-block: $Z = (n), (n-1), (n-2), (n-3), (n-4), (n-5)$

f-block: $Z = (58 \text{ to } 71) \text{ \& } (89 \text{ to } 100)$

d-block: $Z = \text{rest all}$

→ Period: period of next noble gas.

→ Group: $18 + \text{given at. no. } (Z) - \text{atomic no. of next noble gas}$

		Block	Period	Group
eg: (i)	$Z = 12$	s	3	$18 + 12 - 18 = 2$
(ii)	$Z = 9$	p	2	$18 + 9 - 10 = 17$
(iii)	$Z = 21$	d	4	$18 + 21 - 36 = 3$
(iv)	$Z = 33$	p	4	$18 + 33 - 36 = 15$
(v)	$Z = 48$	d	5	$18 + 48 - 54 = 12$
(vi)	$Z = 65$	f	6	3
(vii)	$Z = 118$	p	7	$18 + 118 - 118 = 18$
(viii)	$Z = 58$	f	6	3
(ix)	$Z = 74$	d	6	$18 + 74 - 86 = 6$
(x)	$Z = 81$	p	6	$18 + 81 - 86 = 13$
(xi)	$[Ar] 3d^5 4s^2$	d	4	$2 + 5 = 7$
(xii)	$[Xe] 6s^1$	s	6	1
(xiii)	$[Kr] 4d^{10} 5s^2 5p^2$	p	5	$10 + 2 + 2 = 14$
(xiv)	$[Xe] 4f^{14} 5d^1 6s^2$	f	6	3
(xv)	$[Xe] 5d^{10} 6s^1$	d	6	11

Q:- Calculate maximum no. of elements which can be present in the 9th period of periodic table.

Ans:- Method 1:

Types of orbitals filled in 9th period:

9s ~~5g~~ 6g 7f 8d 9p

↓ ↓ ↓ ↓ ↓

no. of orbitals = $1 + 9 + 7 + 5 + 3 = 25$.

no. of elements = no. of $e^- = 2 \times 25 = 50$.

Method 2:

no. of elements = $2\left(\frac{n+1}{2}\right)^2$ ($\because n$ is odd)

$$= 2\left(\frac{9+1}{2}\right)^2 = 2\left(\frac{10}{2}\right)^2 = 50$$

* Some important points about Modern Periodic Table :

- **Diagonal Relationship**: Due to same Ionic Potential there are three pairs which have similar properties with diagonal element to each other. They are: $(\text{Li} \leftrightarrow \text{Mg})$
 $(\text{Be} \leftrightarrow \text{Al})$
 $(\text{B} \leftrightarrow \text{Si})$
- **Bridging Elements**: Period 3 elements are known as bridging elements as they have valencies similar to subgroup A & subgroup B (acc. to Mendeleev).
- **Typical Elements**: Period 3 elements are known as Typical elements as they represent most of the properties of the elements of their respective periods.
- **Transuranic**: Uranium ($Z=92$) is last naturally occurring element, all elements with atomic no. > 92 are man made, and known as Transuranic Elements.
- **Periodic Table (at Room Temperature) has**:
 - 11 Gaseous Elements: $\text{H}, \text{N}, \text{O}, \text{F}, \text{Cl}, \text{He}, \text{Ne}, \text{Ar}, \text{Kr}, \text{Xe}, \text{Rn}$.
 - 2 liquid Elements: Hg (Metal) + Br (Non Metal).

* Effective Nuclear Charge (Z_{eff} or Z^*):-

- In multielectron species, electrons experience lesser attraction than number of protons present in the nucleus.
- The decrease in force of attraction is due to inner shell electrons, which is called as screening effect or shielding effect.
- Decrease in net force of attraction results in decrease in nuclear charge experienced by outer shell e^- , Net nuclear charge experienced by outer most shell e^- is called effective nuclear charge (Z^*).

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

Where, Z = Nuclear charge.
 Z^* = Effective Nuclear Charge.
 σ = Screening constant.
or shielding constant.

Characteristics of screening constant (σ):

- Screening will be caused by same shell e^- or inner shell electrons only.
- None of outer shell electrons offer screening to inner shell electrons.
- Electrons closer to nucleus offer more shielding, it is known as penetration effect.

Screening offered by e^- in : $K > L > M > N \dots$

* Screening offered by e^- in : $s > p > d > f \dots$
(for a particular shell)

Theoretical calculation of screening constant σ :-

SLATER'S RULE:

① Write electronic configuration in expanded form.

② Rearrange the E.C. as follows:

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

③ (a) If σ on e^- of ns or np is calculated:

(i) σ for each e^- of ns & np (other than 1) is 0.35.

(ii) σ for each e^- of $(n-1)$ shell is 0.85.

(iii) σ for each e^- of $(n-2)$ shell and inner shell is 1.

(iv) σ for each e^- in outer shell of ns and np is 0.

(b) If σ on e^- of nd or nf is calculated:

(i) σ for each e^- of nd or nf (other than 1) is 0.35.

(ii) σ for each e^- on LHS of nd or nf (all inner e^-) is 1.

(iii) σ for each e^- on RHS of nd or nf (all outer e^-) is 0.

eg: Find Z^* on valence e^- of Nitrogen ($Z=7$).

N: $1s^2\ 2s^2\ 2p^3$
 $(1s^2)\ (2s^2\ 2p^3)$

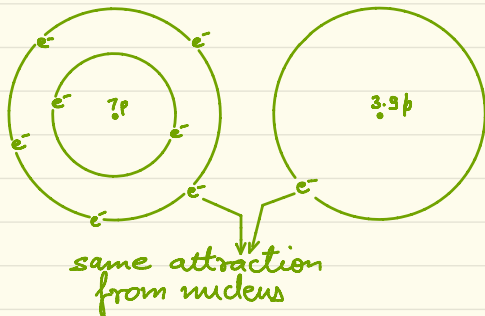
$$\sigma = 2 \times 0.85 + 4 \times 0.35$$

$$\sigma = 1.7 + 1.4$$

$$\sigma = 3.1$$

$$Z^* = (Z - \sigma) = (7 - 3.1)$$

$$Z^* = 3.9$$



Q:- Find Z^* on (i) 4s, (ii) 3d, (iii) 3p, (iv) 3s electron of Fe.

(i) 4s e⁻ of Fe.

$$\begin{aligned}\text{Fe} &: (1s^2)(2s^2 2p^6)(3s^2 3p^6)(3d^6)(4s^2) \\ \sigma &= (10 \times 1) + (14 \times 0.85) + (1 \times 0.35) \\ &= 10 + 11.9 + 0.35 = 22.25 \\ Z^* &= 26 - 22.25 = 3.75\end{aligned}$$

(ii) 3d e⁻ of Fe.

$$\begin{aligned}\text{Fe} &: (1s^2)(2s^2 2p^6)(3s^2 3p^6)(3d^6)(4s^2) \\ \sigma &= (18 \times 1) + (5 \times 0.35) + 0 \\ \sigma &= 18 + 1.75 = 19.75 \\ Z^* &= 26 - 19.75 = 6.25\end{aligned}$$

(iii) 3p e⁻ of Fe.

$$\begin{aligned}\text{Fe} &: (1s^2)(2s^2 2p^6)(3s^2 3p^6)(3d^6)(4s^2) \\ \sigma &= 2 \times 1 + 8 \times 0.85 + 5 \times 0.35 \\ \sigma &= 2 + 6.8 + 1.75 = 10.55 \\ Z^* &= 26 - 10.55 = 15.45\end{aligned}$$

(iv) 3s e⁻ of Fe.

$$\begin{aligned}\text{Fe} &: (1s^2)(2s^2 2p^6)(3s^2 3p^6)(3d^6)(4s^2) \\ \sigma &= 2 \times 1 + 8 \times 0.85 + 5 \times 0.35 \\ \sigma &= 2 + 6.8 + 1.75 = 10.55 \\ Z^* &= 26 - 10.55 = 15.45\end{aligned}$$

Q:- Find Z^* on valence e⁻ of Li, Na, K and Mg.

Li: $(1s^2)(2s^1)$

$$\begin{aligned}\sigma &= 2 \times 0.85 + 0 \times 0.35 = 1.7 \\ Z^* &= 3 - 1.7 = 1.3\end{aligned}$$

K: $(1s^2)(2s^2 2p^6)(3s^2 3p^6)(4s^1)$

$$\begin{aligned}\sigma &= 10 \times 1 + 8 \times 0.85 + 0 = 16.8 \\ Z^* &= 19 - 16.8 = 2.2\end{aligned}$$

Na: $(1s^2)(2s^2 2p^6)(3s^1)$

$$\begin{aligned}\sigma &= 2 \times 1 + 8 \times 0.85 + 0 \times 0.35 = 8.8 \\ Z^* &= 11 - 8.8 = 2.2\end{aligned}$$

Mg: $(1s^2)(2s^2 2p^6)(3s^2)$

$$\begin{aligned}\sigma &= 2 \times 1 + 8 \times 0.85 + 1 \times 0.35 = 9.15 \\ Z^* &= 12 - 9.15 = 2.85\end{aligned}$$

Observations of Slater:

- (i) z^* increases by 0.65 in a period in 2nd & 3rd period.

2nd →	Li 1.3	Be 1.95	B 2.6	C 3.15	N 3.9	O 4.55	F 5.2	Ne 5.85
3rd →	Na 2.2	Mg 2.85	Al 3.5	Si 4.15	P 4.7	S 5.35	Cl 6.0	Ar 6.65

- (ii) z^* increases by 0.9 in a group on moving 2nd to 3rd period.

2nd →	Li 1.3	Be 1.95	B 2.6	C 3.15	N 3.9	O 4.55	F 5.2	Ne 5.85
3rd →	Na 2.2	Mg 2.85	Al 3.5	Si 4.15	P 4.7	S 5.35	Cl 6.0	Ar 6.65

- (iii) z^* remains constant in a group on moving 3rd to 7th period in s-block.

Constant = 2.2	Na 2.2	Mg 2.85	Constant = 2.85
	K 2.2	Ca 1.95	
	Rb 2.2	Sr 2.85	
	Cs 2.2	Ba 2.85	
	Fr 2.2	Ra 2.85	

- (iv) In two e^- species like, He, Li^+ , Be^{2+} etc, screening constant of one e^- for another is 0.30.

eg:- In He ($1s^2$),

$$\sigma = 0.3 \text{ (and not } 0.35)$$

$$z^* = 2 - 0.3 = 1.7 \text{ (and not } 1.65)$$

- (vi) According to Slater, σ and hence z^* on ns and up electron is same.

eg:- z^* on 3s e^- of Fe = z^* on 3p e^- of Fe = 15.45

Practical observation of Effective Nuclear Charge (Z_{eff})

→ In a period, on moving from left to right proton increases by 1 and effect of proton is 100%, while no. of electron increase by 1 but effect of electron is less than 100%, (since e^- is movable).
(proton (p) \uparrow by 1 \rightarrow effect of $p = 100\%$)
(electron (e^-) \uparrow by 1 \rightarrow effect of $e^- < 100\%$)

So, overall as atomic no. (Z) \uparrow $Z_{eff} \uparrow$.

Also, number of shell remain same and $Z_{eff} \uparrow$
so, force of attraction (F_A) \uparrow (left to right).

→ In a group, On moving top to bottom:

a) In s-block, $Z_{eff} \uparrow$ slightly.

So, \uparrow in shell no. (n) dominates small \uparrow in Z_{eff} .
hence, force of attraction (F_A) \downarrow down the grp.

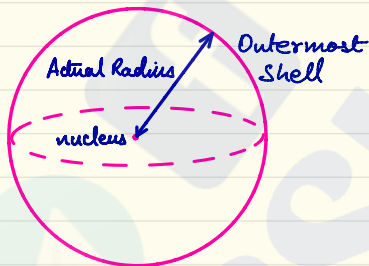
b) In d-block, $Z_{eff} \uparrow$ rapidly, due to poor shielding d & f e^- .
So, large \uparrow Z_{eff} dominates \uparrow in shell no. (n)
hence force of attraction (F_A) \uparrow down the grp.

c) In p-block, initially $Z_{eff} \uparrow$ rapidly due to less no. of p-electrons and large no. of d & f electrons.
So, large \uparrow in Z_{eff} dominates \uparrow in shell no. (n)
hence, $F_A \uparrow$ down the grp, initially in p-block.

But later from (p' to p'') in p-block, no. p-electrons \uparrow and no. of d & f-electrons remains same, so $Z_{eff} \uparrow$ slowly.
So, \uparrow in shell no. (n) dominates small \uparrow in Z_{eff} .
hence, $F_A \downarrow$ down the grp, later in p-block.

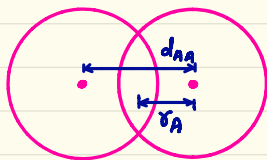
* Periodic trends in Property of Elements:

(I) ATOMIC SIZE : Theoretically, atomic size is the shortest distance between nucleus and outermost electron of isolated gaseous atom.



- It is difficult to measure the atomic size practically, because:
- (i) Existence of isolated atom is difficult.
 - (ii) Position of moving electron cannot be predicted accurately. (Heisenberg's uncertainty principle).
- Therefore, determination of atomic size is done by calculating half of the internuclear distance between 2 atoms of an element.
- Based on the Nature of the element and Chemical bonds present with in them, different type of radius are defined :
- 1) Covalent radius (For Non metals)
 - 2) Metallic radius (for Metals)
 - 3) Vanderwaal's radius (For Noble gases)

1. Covalent radius: Covalent radius is half of internuclear distance between two singly bonded homoatoms



$$r_A = \frac{d_{AA}}{2}$$

→ Three types of covalent radii can be given between two homoatoms

- Single bonded covalent radii (SBCR) = half of singly bonded homoatom
- Double bonded covalent radii (DBCR) = half of doubly bonded homoatom
- Triple bonded covalent radii (TBCR) = half of triply bonded homoatom

eg: distance between C-C > dist. between C=C > dist. between C≡C

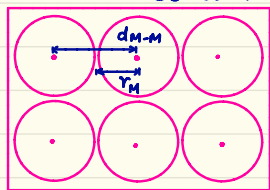
$$d_{C-C} > d_{C=C} > d_{C\equiv C}$$

$$\frac{d_{C-C}}{2} > \frac{d_{C=C}}{2} > \frac{d_{C\equiv C}}{2}$$

$$\text{SBCR of carbon} > \text{DBCR of carbon} > \text{TBCR of carbon}$$

→ By default, Single bonded covalent radius (SBCR) is covalent radius, because atoms of non-metals are bonded at least by single covalent bond, while they may or may not form, double / tripple bonds.

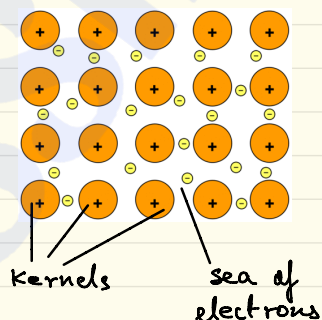
2. Metallic Radius : Half of the internuclear distance between two adjacent atoms of a metal bonded by metallic bond in crystalline lattice structure.



$$r_M = \frac{d_{M-M}}{2}$$

Electron Sea Model : Metallic Bond:

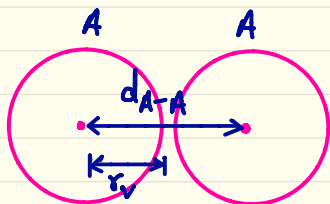
→ In early 1900's, Paul Drude came up with "sea of electrons" metallic bonding theory by modelling metals as mixture of kernels (positive nucleus + inner shell electrons) and outer most shell electrons.



→ An array of positively charged Kernels is present in a sea of freely moving e^- , in gaps of these Kernels.

→ The negative charge of e^- sea helps to bind kernels in the array.

3. Vanderwaal's Radius : Half of internuclear distance between two adjacent non bonded atoms of an element in a highly compressed state.



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

→ In covalent bond, electrons are concentrated between the nucleus while e^- in metallic bonds are movable so attraction in covalent bond > Metallic bond.

Also, vanderwaal forces are weakest, so vander-waal radius is largest.

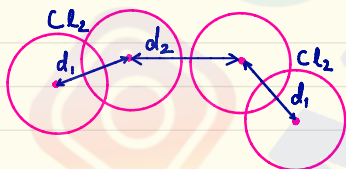
So, for a particular atom:

$$\text{Covalent radius} < \text{Metallic radius} \ll \text{Vanderwaal radius}$$

→ for noble gas, only vanderwaal radius is calculated.

→ for non metals, covalent radius and vander waal radius, both can be calculated.

eg: for chlorine,

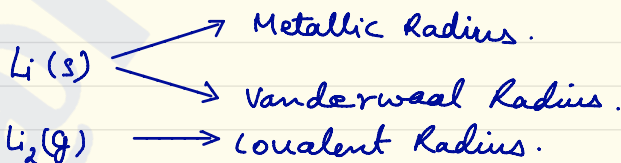


Covalent radius of $\text{Cl} = \frac{d_1}{2}$

Vanderwaal radius of $\text{Cl} = \frac{d_2}{2}$

→ for metals, vanderwaal radius, covalent radius and metallic radius, all three can be calculated.

eg: for Lithium:

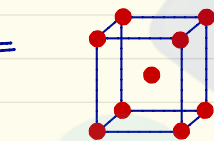


Element	Radius that may be calculated	Default Radius
Metals	Metallic, covalent, vanderwaal	Metallic
Non Metals	Covalent, vanderwaal	Covalent
Noble gases	vanderwaal	vanderwaal

Factors affecting Atomic Size:

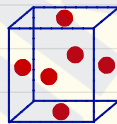
- (i) Nuclear charge (Z) = $Z \uparrow$ attraction on $e^- \uparrow$ Size \downarrow
- (ii) Shielding (σ) = $\sigma \uparrow$ attraction on $e^- \downarrow$ Size \uparrow
- (iii) Effective nuclear charge (Z^*) = $Z^* \uparrow$ attraction on $e^- \uparrow$ size \downarrow
- (iv) Shell no. (n) = $n \uparrow$ size of atom \uparrow

(v) Crystal Lattice =



Body Centered Cubic

>



Face Centered Cubic

Periodic Variation of atomic size:

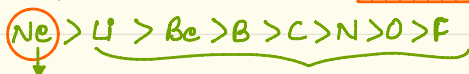
(i) In s-block and p-block:

Along the period:

→ On moving left to right shell no. (n) remains same, but $Z^* \uparrow$ therefore atomic size \downarrow

→ Since Vanderwaal radius exists for noble gas, they are largest in their respective periods. It is also explained by excessive shielding caused by fully filled orbitals.

eg: size of elements of 2nd period



Vanderwaal
radius

Covalent ($Z^* \uparrow$ Size \downarrow)
radius

eg: Order of Vanderwaal radius of elements of 2nd period

order of size: $Li > Be > B > C > N > O > \boxed{F < Ne}$

$\textcircled{Ne} > Li > Be > B > C > N > O > F$

Vanderwaal
radius

Covalent ($Z^* \uparrow$ Size \downarrow)
radius

When nothing is mentioned size of elements is compared on the basis of default radius.

Order of Vanderwaal radius: $Li > Be > B > C > N > O > F > \textcircled{Ne}$
($Z^* \uparrow$ Size \downarrow)

When radius of elements is compared on same scale then, Z^* is only factor in the period and shell no. n is only factor in group.

→ For Period 1, 2, 3: Size of group 18 > group 1 > group 2...
 $H < He$
 $Li < Ne$
 $Na < Ar$

→ For Period 4, 5, 6, 7: Size of group 1 > group 18 > group 2...
 $K > Kr$
 $Rb > Xe$
 $Cs > Rn$
 $Fr > Og$

→ exception: $Pb > Bi$ (expected) ✗
 $\textcircled{Pb < Bi}$ (actual) ✓
due to crystal lattice.

Along the group:

→ On moving top to bottom shell no. (n) ↑ so size ↑

→ Although the size increases, the rate of increase of size, decreases down the group, due to increase in no. of poor shielding e^-

exception: In group 13 order of size is:

expected: $B < Al < Ga < In < Tl$ ✗

Actual: $B < Al > Ga < In < Tl$ ✓

$B < Ga < Al < In < Tl$ ✓

↓
due to 10 poor shielding 3d electrons in Ga, $Z^* \uparrow$. Increase in Z^* dominates over ↑ in shell no.

→ In group 14 order of size is: $C < Si < Ge < Sn < Pb$
(no exception) shell no. (n) ↑ so size ↑

→ top to bottom increment in size ↓

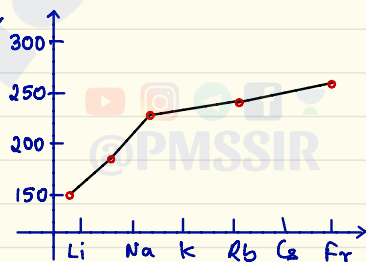
Li Na K Rb Cs Fr

x

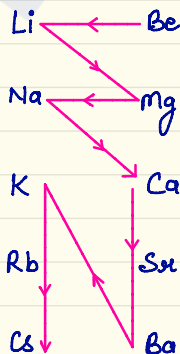
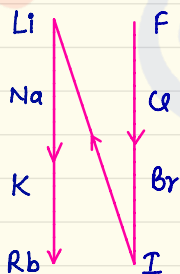
y

z

diff in size $x > y > z$



→ Data based orders:

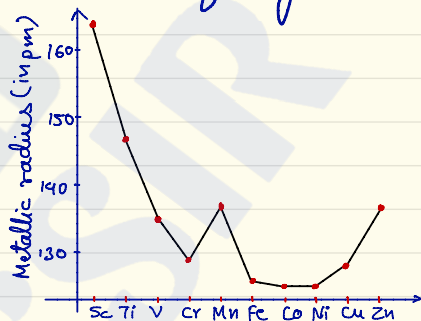
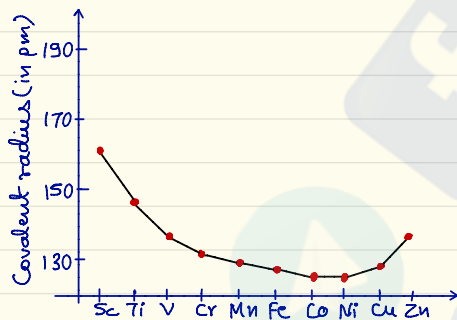


Small → large

(ii) In d-block:

→ Along the period:

On moving left to right size first decreases, then remains constant in middle and then finally increases



→ Factor 1: $z^* \uparrow$

Factor 2: no. of unpaired e-
M-M bonding
Size

Overall order of size from 1 & 2

Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn
	Size ↑				Size ↑			Size ↑	
	↑				↓			↓	
	↑				↓			↓	
	Size ↓				Size ↑			Size ↑	
	Size ↓				Size almost constant			Size ↑	

→ Along the group:

	3	4	5	6	7	8	9	10	11	12
3d										
4d										
5d										

$3d < 4d < 5d$
 $n \uparrow$ Size ↑

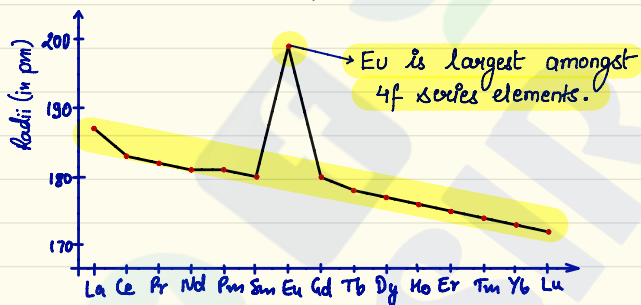
$3d < 4d \approx 5d$ or $3d < 4d > 5d$

Lanthanide Contraction

→ Lanthanide Contraction: In d-block on 4d to 5d series after Lanthanum size decreases or remains almost same due to increase in z^* of 14 poor shielding 4f electrons, this is known as Lanthanoid contraction.

(iii) In f-block :

→ On moving left to right size decreases (generally), due to increase in nuclear charge (Z).

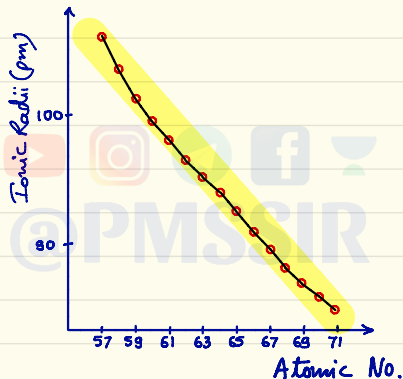
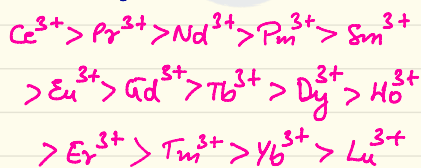


→ On moving left to right, the ↓ in size of Lanthanoids is less than ↓ in size of Actinoids, because no. of poor shielding electrons are more in 7th period.

→ Actinide contraction: decrease in size of elements in 7th period due to poor shielding of 5f and 6d electrons.

→ Actinide contraction is higher than Lanthanoid contraction. due to presence of 5f electrons in actinoids.

→ Size of M^{3+} ↓ regularly from left to right in 4f series.



Ionic Radius :

→ It is of 2 types ① Cationic Radius and ② Anionic Radius

① Cationic Radius :

When neutral atom loses electron it forms cation.
Cation is always smaller than neutral atom.

As positive charge on cation increases size of cation decreases

	A	→	A ⁺	→	A ²⁺	→	A ³⁺
no. of protons	Z		Z		Z		Z
no. of electrons	Z		Z-1		Z-2		Z-3

→
Z remains same, σ ↓, Z^* ↑, size ↓

② Anionic Radius :

When neutral atom gains electron it forms anion.
Anion is always larger than neutral atom.

As no of negative charge on anion increases, size of anion inc.

	A	→	A ⁻	→	A ²⁻	→	A ³⁻
no. of protons	Z		Z		Z		Z
no. of electrons	Z		Z+1		Z+2		Z+3

→
Z remains same, σ ↑, Z^* ↓, size ↑

③ Size of Isoelectronic species: It is compared by (Z/e) ratio
(Z/e) ratio ↑ Ionic radius ↓

eg: Na^+ , Mg^{2+} , F^- , N^{3-} , Al^{3+}

Z → 11 12 9 7 13

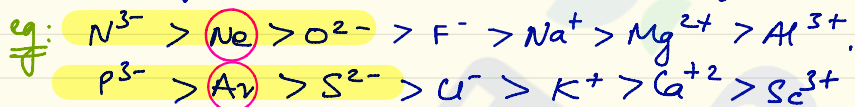
e^- → 10 10 10 10 10

Z/e → 1.1 1.2 0.9 0.7 1.3

Z/e order → $\text{Al}^{3+} > \text{Mg}^{2+} > \text{Na}^+ > \text{F}^- > \text{N}^{3-}$

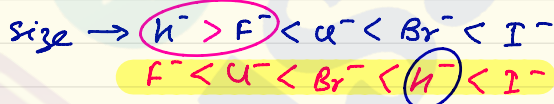
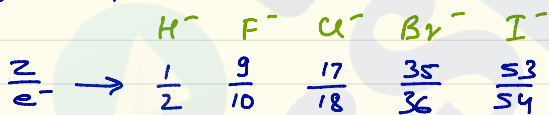
Ionic radius → $\text{Al}^{3+} < \text{Mg}^{2+} < \text{Na}^+ < \text{F}^- < \text{N}^{3-}$

→ Although noble gas are isoelectronic with others, but z/e can't be used because of different type of radius, ionic radius for ions and vanderwaal radius for noble gases.

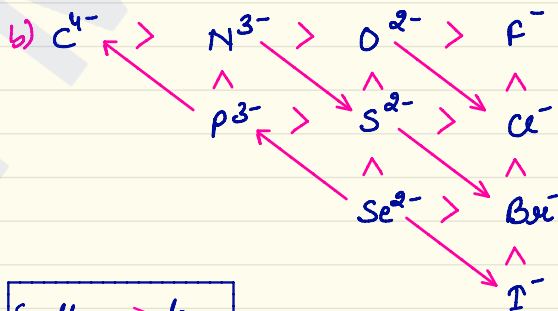
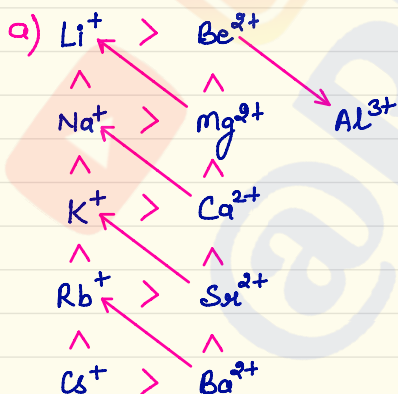


→ Smallest cation = H^+

Smallest anion = F^- (not H^-)



④ Some ionic radius based on data :

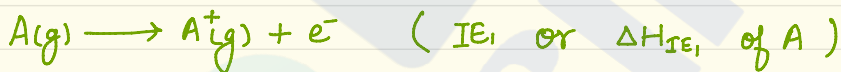


Small \rightarrow large

(II) Ionisation Energy and Ionisation Enthalpy

→ Ionisation Energy: It is the amount of energy required to remove most loosely held electron from isolated gaseous atom/ion to form another isolated gaseous ion.

→ Ionisation Enthalpy (ΔH_{IE}): Change in enthalpy when an electron is removed from outermost shell of isolated gaseous atom/ion.

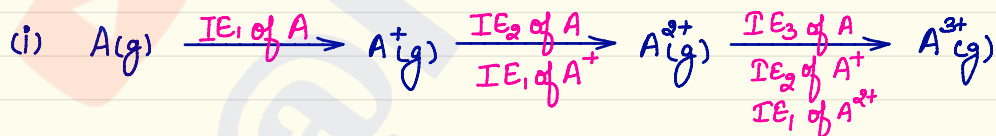


→ By default all ionisation energy [IE] values are specified at zero Kelvin (0 K) hence at any other temperature, change in energy for ionisation process is given by ionisation enthalpy (ΔH_{IE}).

$$\Delta H_{IE} = IE + \frac{5}{2} RT$$

$$\Delta H_{IE} \text{ (at zero Kelvin)} = IE$$

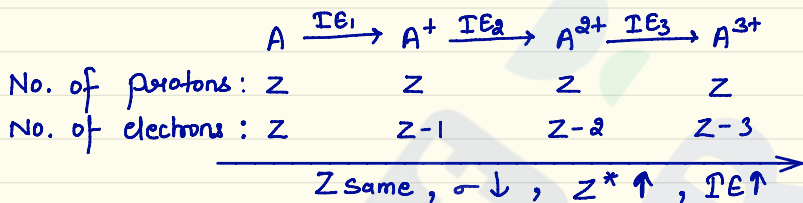
Some Important points about IE/ ΔH_{IE} :



(iii) Ionisation is always endothermic process (energy is required) for neutral atom or cation.

(iv) Successive ionisation energy is always increasing

$$IE_1 < IE_2 < IE_3 < IE_4 < \dots$$



(v) It is measured in eV/atom or Kcal/mol or KJ/mol

$$1 \text{ eV/atom} = 96.4 \text{ KJ/mol}$$

$$4.2 \text{ KJ/mol} = 1 \text{ Kcal/mol}$$

Factors Affecting Ionisation Energy/ Ionisation Enthalpy:

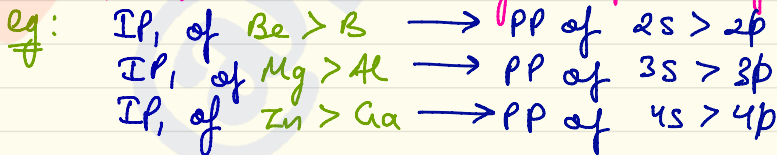
(i) Effective nuclear charge (Z^*) \uparrow attraction \uparrow $IE \uparrow$

(ii) Screening constant (σ) \uparrow attraction \downarrow $IE \downarrow$

(iii) Size/shell no. (n) \uparrow attraction \downarrow $IE \downarrow$

(iv) Penetration power of subshells \uparrow attraction \uparrow $IE \uparrow$.

Penetration Power (P.P.) of $s > p > d > f$.



(v) Extra stable electronic configuration, half filled (p^3, d^5, f^7) and fully filled (p^6, d^{10}, f^{14}) electronic configuration require slightly higher ionisation energy.

Periodic Variations of IE / ΔH_{IE}

→ Periodic variations of ionisation energy (IE), ionisation enthalpy (ΔH_{IE}) and ionisation potential (IP) are same.

(i) In s-block and p-block:

→ In a period, on moving left to right $IE \uparrow$ (generally)
(extra stable electronic configurations like ns^2 or $ns^2 np^3$ or $ns^2 np^6$ have higher IE than next element).

eg: Order of IE_1 of second period elements:

expected: $Li < Be < B < C < N < O < F < Ne$

actual: $Li < B < Be < C < O < N < F < Ne$

eg: Order of IP_1 of s-block and p-block elements:

period 3: $Na < Al < Mg < Si < S < P < Cl < Ar$

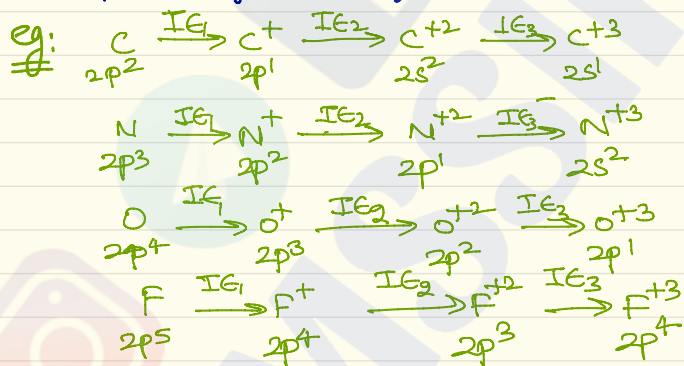
period 4: $K < Ga < Ca < Ge < Se < As < Br < Kr$

period 5: $Rb < Sr < In < Sn < Sb < Te < I < Xe$

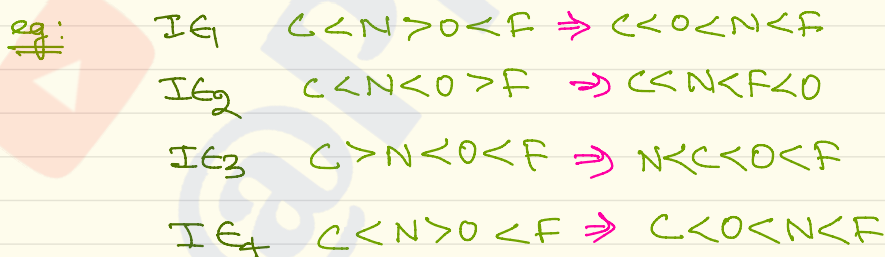
period 6: $Cs < Ba < Tl < Bi < Pb < Po < At < Rn$

due to size

→ Comparison of n^{th} ionization energy
 for IE_1 Configuration of A^0 is compared
 for IE_2 Configuration of A^+ is compared
 for IE_3 Configuration of A^{2+} is compared
 for IE_n Configuration of A^{n-1} is compared

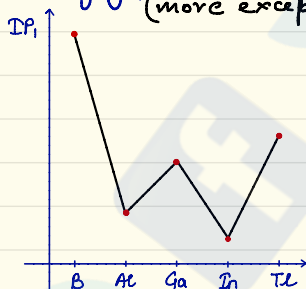


→ Always remove $(n-1)e^-$ in Comparison of IE_n .



→ In a group, On moving top to bottom $n \uparrow$ so $IE \downarrow$ (generally)

exceptions: (i) IP_1 of group 13: (more exception)



expected:

$B > Al > Ga > In > Tl$ ✗

Actual:

$B > Al < Ga > In < Tl$ ✓

overall:

$B > Tl > Ga > Al > In$ ✓

(ii) IP_1 of group 14: $C > Si > Ge > Sn < Pb$
(less exception) $C > Si > Ge > Pb > Sn$

(iii) IP_1 of group 15: $N > P > As > Sb > Bi$
(no exception)

(iv) IP_1 of $Sn < Pb$

(v) IP_1 of $Cs \approx Fr$

(vi) IP_1 of $Ba < Ra$

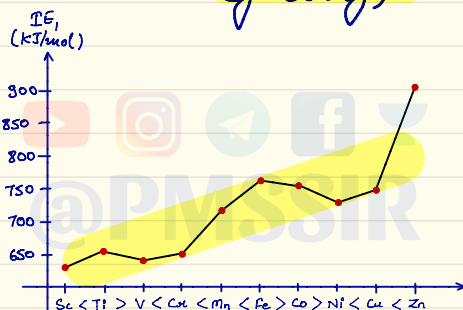
(ii) In d-block:

→ In a Period, on moving from left to right $Z^* \uparrow$ $IE \uparrow$ (generally)

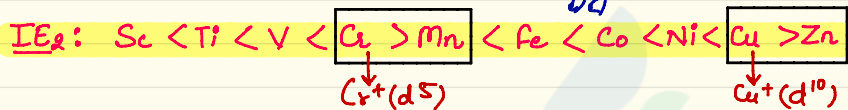
→ When an e^- is removed from (ns), positive charge appears, hence second (ns) e^- jumps to $(n-1)d$. If no. of exchange increases sufficiently, decrement in IE observed.

IE_1 : $Sc < Ti > V < Cr < Mn$

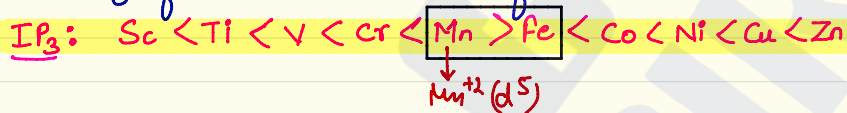
$< Fe > Co > Ni < Cu < Zn$



→ IE_2 from left to right increases but Cr and Cu have higher value of IE_2 value than Mn and Zn respectively due to more stable electronic configuration.



→ IE_3 of elements of 3d-series increases from left to right, but IP_3 of Mn > Fe because of extra stable E.C.



→ In a group, On moving top to bottom for :

IP_1 of group 3: 3d > 4d > 5d > 6d → Sc > Y > La > Ac

IP_1 of group 10: 5d > 4d > 3d → Pt > Pd > Ni

IP_1 of group 8, 9, 11, 12: 5d > 3d > 4d → Os > Fe > Ru

Ir > Co > Rh

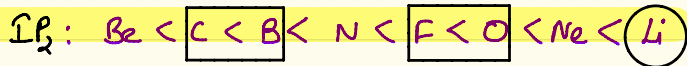
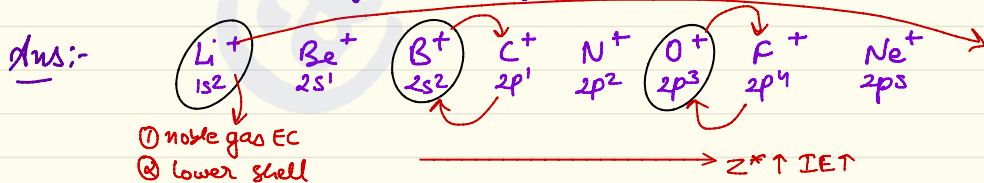
Au > Cu > Ag

Hg > Zn > Cd

(iii) In f-block :

→ On moving left to right size ↓ IE ↑ (generally), due to increase in nuclear charge (Z).

Q:- Give order of IP_2 of 2nd period element?



Application of IE / ΔH_{IE}

(i) Metallic Character $\propto \frac{1}{IE}$

that is down the group metallic character \uparrow
and left to right in a period metallic character \downarrow

(ii) Reducing Nature $\propto \frac{1}{IE}$

Reducing Nature of group 1 elements (in gas): $Li < Na < K < Rb < Cs$

(iii) Elements with lower IE have greater tendency to form cation.

(iv) To determine the No. of valence e^- of representative elements:

No. of valence electron = No. of ionisation before a sudden jump in successive IEs

eg: Successive IE of an element: 14, 16, 20, 24, 75, 100
No of valence electron = 4 (Sudden jump between IE_3 & IE_4)

(v) Stability of oxidation states of an element:

Depending upon the difference in successive IE stable oxidation state can be predicted.

$[IE_n - IE_{(n-1)}] > 16 \text{ eV/atom}$: lower O.S more stable

$[IE_n - IE_{(n-1)}] < 11 \text{ eV/atom}$: higher O.S more stable

eg: $Al \rightarrow Al^+$
 $Al^+ \rightarrow Al^{2+}$ } $\Delta IE (IE_2 - IE_1) = 12.8 \text{ eV/atom} > 16 \text{ eV/atom} \rightarrow \text{lower O.S stable}$
So Al^+ is more stable than Al^{2+}

$Al^+ \rightarrow Al^{2+}$
 $Al^{2+} \rightarrow Al^{3+}$ } $\Delta IE (IE_3 - IE_2) = 6 \text{ eV/atom} < 11 \text{ eV/atom} \rightarrow \text{higher O.S stable}$
So Al^{3+} is more stable than Al^{2+}

Overall Order of stability: $Al^{3+} > Al^+ > Al^{2+}$

(vi) Inert Pair Effect

→ In p-block, on moving down the group stability of higher oxidation state ↓ and stability of lower oxidation state ↑.

→ Because, as we move down the group, no. of poor shielding d & f electrons ↑, so attraction on (ns) electrons ↑ which makes (ns) e⁻, inert for bond formation.

grp 13	grp 14	grp 15	
B → +3	C → +4	N → +5	Stability of Higher O.S. > Lower O.S.
Al → +1 → +3	Si → +2 → +4	P → +3 → +5	
Ga → +1 → +3	Ge → +2 → +4	As → +3 → +5	
In → +1 → +3	Sn → +2 → +4	Sb → +3 → +5	
Tl → +1 → +3	Pb → +2 → +4	Bi → +3 → +5	
			lower > Higher O.S. O.S.

Down the Group
Stability of
Higher O.S. ↓
Lower O.S. ↑

→ Order of stability:

- i) grp 13: $\text{Ga}^{+1} < \text{In}^{+1} < \text{Tl}^{+1}$
 $\text{Ga}^{+3} > \text{In}^{+3} > \text{Tl}^{+3}$
- ii) grp 14: $\text{Ge}^{+2} < \text{Sn}^{+2} < \text{Pb}^{+2}$
 $\text{Ge}^{+4} > \text{Sn}^{+4} > \text{Pb}^{+4}$
- iii) grp 15: $\text{As}^{+3} < \text{Sb}^{+3} < \text{Bi}^{+3}$
 $\text{As}^{+5} > \text{Sb}^{+5} > \text{Bi}^{+5}$
- iv) $\left. \begin{array}{l} \text{Pb}^{+2} > \text{Pb}^{+4} \\ \text{Bi}^{+3} > \text{Bi}^{+5} \end{array} \right\} \text{Pb}^{+4}, \text{Bi}^{+5} \text{ are good oxidants}$
- v) $\text{Sn}^{+2} < \text{Sn}^{+4} \rightarrow \text{Sn}^{+2} \text{ is good reductant}$

Down the Group
Stability of
Higher O.S. ↓
Lower O.S. ↑

(III) Electron Affinity / Electron Gain Enthalpy (ΔH_{eg}):

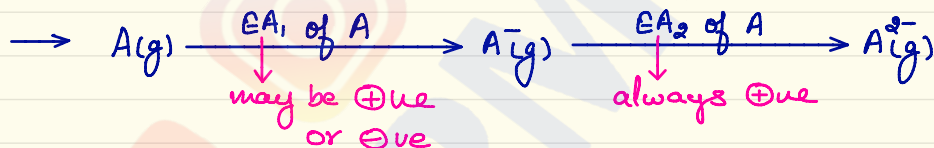
→ Electron Affinity (EA): It is the amount of energy released when an electron is added to an isolated gaseous atom or ion

→ Electron gain Enthalpy (ΔH_{eg}): Change in enthalpy when an e^- is added to isolated gaseous atom or ion

→ By default all electron affinity values are given at zero Kelvin (0 K), but at other temperatures electron gain enthalpy is given.

$$\Delta H_{eg} = -EA - \frac{5}{2}RT$$

$$\Delta H_{eg} \text{ (at zero Kelvin)} = -EA$$



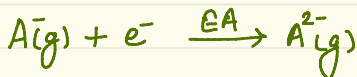
ΔH_{eg_1} may be \ominus ve or \oplus ve for neutral atom

ΔH_{eg_2} always \ominus ve for neutral atom

eg: In the process,

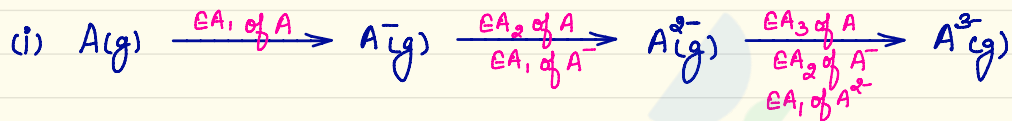


If energy released is 20 eV/atom
then $EA = 20 \text{ eV/atom}$
and $\Delta H_{eg} = -20 \text{ eV/atom}$



If energy absorbed is 30 eV/atom
then $EA = -30 \text{ eV/atom}$
and $\Delta H_{eg} = 30 \text{ eV/atom}$

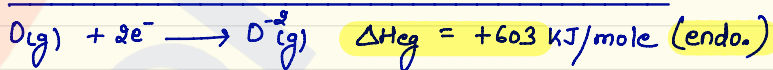
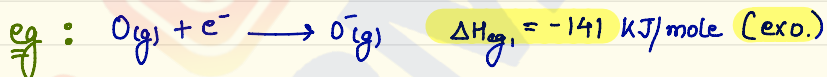
Some Important points about EA/ ΔH_{eg} :



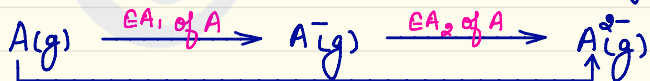
(iii) Magnitude of successive EA of neutral atom is \uparrow .
 $|EA_1| < |EA_2| < |EA_3| < \dots$

(iv) Generally first electron addition of an isolated gaseous atom is an exothermic process.
 (except for stable electronic configurations)

(v) Second electron addition of an isolated gaseous atom is always endothermic process due to interelectronic repulsions



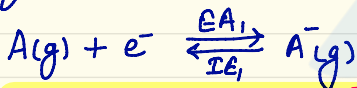
(vi) Formation of polynegative anions (O^{2-} , N^{3-} etc) is always endothermic



$$EA_1 + EA_2 = \text{always } \ominus \text{ve}$$

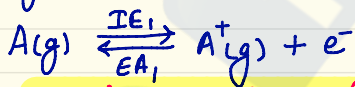
Relation between Electron affinity and Ionisation energy:

(i) Electron affinity of neutral atom = Ionisation energy of its anion



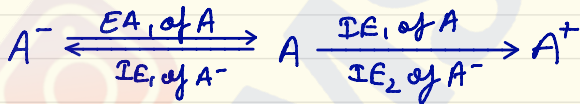
$$EA_1 \text{ of } A = IE_1 \text{ of } A^- \longrightarrow (1)$$

(ii) Ionisation energy of neutral atom = Electron affinity of its cation



$$IE_1 \text{ of } A = EA_1 \text{ of } A^+ \longrightarrow (2)$$

(iii) Ionisation energy of neutral atom is always greater than electron affinity of neutral atom



\therefore Successive IE is always increasing
so $IE_1 \text{ of } A^- < IE_2 \text{ of } A^-$

$$EA_1 \text{ of } A < IE_1 \text{ of } A \longrightarrow (3)$$

	EA	ΔH_{eg}	IE	ΔH_{ie}
if energy is released	⊕ve	⊖ve	⊖ve	⊖ve
if energy is absorbed	⊖ve	⊕ve	⊕ve	⊕ve

Factors Affecting Electron Affinity:

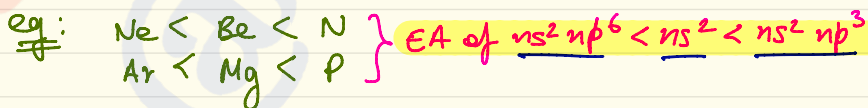
- (i) Effective nuclear charge (Z^*) \uparrow attraction \uparrow EA \uparrow .
- (ii) Screening constant (σ) \uparrow attraction \downarrow EA \downarrow .
- (iii) Shell no. (n) \uparrow attraction \downarrow EA \downarrow .
- (iv) Electron affinity of elements having extra stable electronic configuration (half filled and fully filled) is very less or 0ve. for such elements electron gain enthalpy (ΔH_{eg}) will be positive

s^2		p^3	p^6
Be		N	He
Mg		P	Ne
Ca		As	Ar
Sr		Sb	Kr
Ba			Xe
			Rn

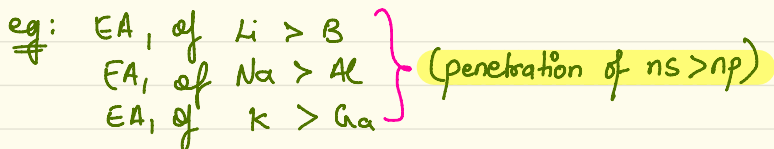
EA is 0ve

In these groups EA \uparrow down the group

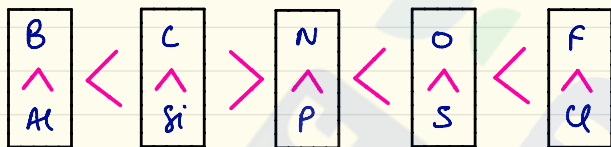
Stability of half filled $<$ fully filled ($ns^2 < ns^2 np^6$)
 so EA of half filled $>$ fully filled ($ns^2 > ns^2 np^6$)



(v) Penetration power \uparrow EA \uparrow .



(vi) Electron density of 2nd period > 3rd period elements
 so, addition of electrons to 2nd period elements is difficult, hence EA of 2nd period element is less than 3rd period element (in p-block).



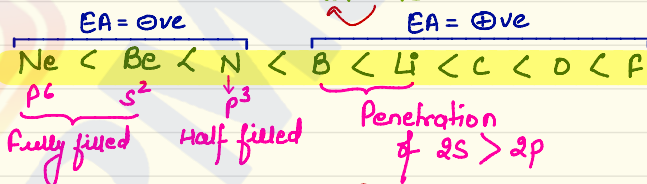
eg: EA₁ of O < S < F < Cl
 EA₁ of N < P < O < S

Periodic variations of electron affinity:

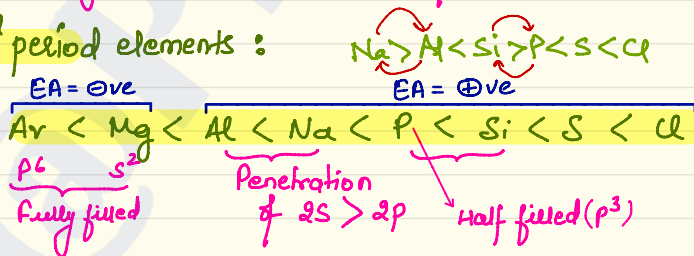
In s-block and p-block:

→ In a period,

→ EA of 2nd period elements:



→ EA of 3rd period elements:



→ In a group,

→ EA of grp 17: $\boxed{\text{F} < \text{Cl}} > \text{Br} > \text{I} \Rightarrow \boxed{\text{Cl} > \text{F}} > \text{Br} > \text{I}$

→ EA of grp 16: $\boxed{\text{O} < \text{S}} > \text{Se} > \text{Te} \Rightarrow \text{S} > \text{Se} > \text{Te} > \boxed{\text{O}}$

Q:- Compare IE of F, F⁻, Cl, Cl⁻.

Ans:-

- ① IE of F > Cl \Rightarrow down the group IE \downarrow .
 - ② IE of F > F⁻
 - ③ IE of Cl > Cl⁻
 - ④ IE of F⁻ > Cl⁻ \Rightarrow EA of F < Cl.
- } Successive IE is higher

from ① to ④ \Rightarrow IE of F > Cl > Cl⁻ > F⁻

(III) Electronegativity:

\rightarrow The tendency of an element to attract shared pair of e⁻ in a covalent bond is called electronegativity.

\rightarrow It is not a measurable quantity. However, a relative numerical value is given on various scales.



$$EN_A = EN_A$$

$$\Delta EN = 0$$

Non polar bond



$$EN_A > EN_B$$

$$\Delta EN \neq 0$$

Polar bond

\rightarrow Electronegativity of some elements on Pauling's scale are:

Li 1.0	Be 1.5	B 2.0	C 2.5	N 3.0	O 3.5	F 4.0
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Na 0.9	Mg 1.2	Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0
-----------	-----------	-----------	-----------	----------	----------	-----------

K
0.8

Rb
0.8

Cs
0.7

Practically EN of :

P > H

N > Cl

S > C

Br
2.8

I
2.6

Factors affecting Electronegativity:

- (i) Nuclear charge (Z) \uparrow attraction \uparrow EN \uparrow
- (ii) Screening constant (σ) \uparrow attraction \downarrow EN \downarrow
- (iii) Effective nuclear charge (Z^*) \uparrow attraction \uparrow EN \uparrow
- (iv) Shell no (n) \uparrow attraction \downarrow EN \downarrow
- (v) Oxidation state \uparrow attraction \uparrow EN \uparrow
eg: $\text{Fe}^{3+} > \text{Fe}^{2+}$
- (vi) %s-character \uparrow attraction \uparrow EN \uparrow
eg: EN of sp^3 hybridised C $<$ sp^2 hyb. C $<$ sp hyb. C
- (vii) EN of surrounding atom \uparrow , EN of central atom \uparrow
eg: EN of C in $\text{CH}_4 < \text{CH}_3\text{I} < \text{CH}_3\text{Cl} < \text{CH}_3\text{F}$
(\because EN of $\text{F} > \text{Cl} > \text{I} > \text{H}$)
- (viii) No. of more electronegative surrounding atom \uparrow EN of central atom \uparrow .
eg: EN of C in $\text{CH}_4 < \text{CH}_3\text{F} < \text{CH}_2\text{F}_2 < \text{CHF}_3$

Calculation of EN of elements:

- (i) Pauling's method (On Pauling's scale): Based on bond energies

$$|X_A - X_B| = 0.2081 \sqrt{E_{A-B} - \sqrt{E_{A-A} \times E_{B-B}}} = 0.208 \sqrt{\Delta}$$

where, E_{A-A} = Bond energy of pure covalent bond of A-A in Kcal/mole
 E_{B-B} = Bond energy of pure covalent bond of B-B in Kcal/mole
 E_{A-B} = Bond energy of pure covalent bond of A-B in Kcal/mole
 X_A and X_B are EN of A and B respectively
 Δ = Ionic resonance energy.

(ii) Mulliken's method (On Mulliken's scale):

→ EN of element A on Mulliken scale = $\frac{IE \text{ of A} + EA \text{ of A}}{2}$

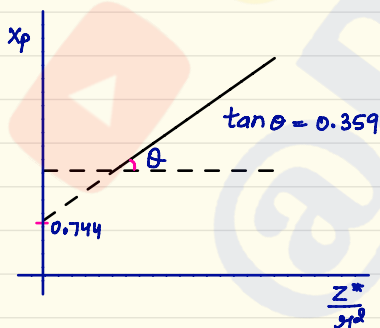
$$X_M = \frac{IE + EA}{2} \quad \text{IE and EA are in eV/atom}$$

→ Numerically EN of element on Mulliken scale is larger value than EN of that element on Pauling Scale.

$$\rightarrow X_p \approx \frac{X_M}{2.8} = \frac{IE + EA}{5.6}$$

Where, X_p = EN of A on Pauling Scale
 X_M = EN of A on Mulliken Scale

(iii) Allred-Roschov's Method (On Pauling's Scale):



$$X_p = 0.359 \frac{Z^*}{r^2} + 0.744$$

Where,

X_p = EN of A on Pauling's scale
 Z^* = Effective Nuclear charge
 r = Covalent radius of A (in Å)

Applications of Electronegativity

(i) Metallic and Non Metallic Nature:

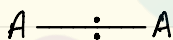
Generally metals have low EN and non metals have high EN.
So, $\text{non-metallic nature} \propto \text{EN}$

(ii) Polarity of Bond $\propto \text{EN}$

Covalent Bond

Polar

e^- cloud is unsymmetrically distributed between bonded atoms.



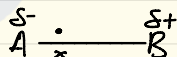
$$E_A = E_A$$

$$\Delta EN = 0$$

Non polar bond

Non-Polar

e^- cloud is symmetrically distributed between bonded atoms.



$$E_A > E_B$$

$$\Delta EN \neq 0$$

Polar bond

(iii) Hanney - Smith Equation:

$$\% \text{ Ionic Character} = 16 |X_A - X_B| + 3.5 (X_A - X_B)^2$$

as $(X_A - X_B) \uparrow$ % Ionic Character in Covalent bond \uparrow

if $\Delta EN = 2.1$ then Ionic Character = Covalent Character

if $\Delta EN > 2.1$ then AB is Ionic Bond

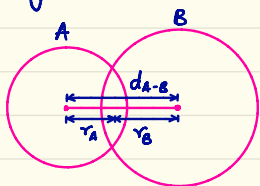
if $\Delta EN < 2.1$ then AB is Covalent

(iv) Bond Energy $\propto \Delta EN$

as $\Delta EN \uparrow$ % I.C. \uparrow Bond energy \uparrow

(v) Bond length (Shoemaker and Stevenson Eqⁿ):

→ Theoretically: for heteroatomic molecule



$$d_{A-B} = r_A + r_B$$

$$\text{also, } r_A \neq r_B \neq \frac{d_{A-B}}{2}$$

→ Practically, due to ΔEN , partial charge generates on atoms and extra attraction is observed

So, Observed $d_{A-B} < \text{Theoretical } d_{A-B}$

→ Shoemaker and Stevenson Eqⁿ

$$d_{A-B} = r_A + r_B - 0.09 |x_A - x_B|$$

where, d_{A-B} = Internuclear distance between A and B
or bond length of A-B.

r_A = covalent radius of A

r_B = covalent radius of B

→ Bond length $\propto \frac{1}{\Delta EN}$

(vi) Nomenclature of Inorganic compounds:

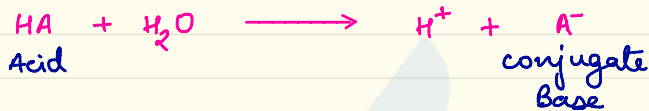
→ While writing names of covalent compounds, names of atoms with lower EN are written first followed by names of atoms with higher EN

eg: Cl_2O → dichlorooxide ✓
Oxodichloride ✗

→ While writing molecular formula of compound, atoms with lower EN are written on left and atoms with higher EN are written on right

eg: Oxygendifluoride → F_2O ✗
 OF_2 ✓

(vii) Nature of Hydride (H-A):



→ stability of conjugate base (A^-) \uparrow , tendency to loose H^+ \uparrow , hence, Acidic nature of acid (HA) \uparrow .

→ In a group

Stability of $\text{A}^- \propto$ Size of A \propto Acidic nature of HA

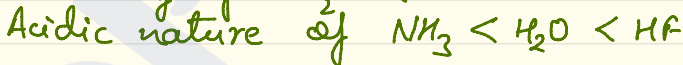
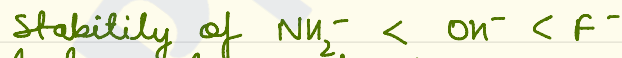
→ In a period

Stability of $\text{A}^- \propto$ EN of A \propto Acidic nature of HA

eg:- Acidic nature of NH_3 , H_2O , HF .



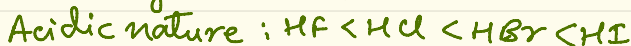
In period, on moving left to right size remains almost same due to same no. of shell, But EN of atom \uparrow so, stability of \ominus charge \uparrow hence acidic nature \uparrow .



eg:- Acidic nature of HF , HCl , HBr , HI .



down the grp, size \uparrow ,
dispersion of \ominus charge
 \uparrow stability of conjugate \uparrow
acidic nature of acid \uparrow



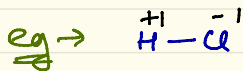
(viii) Oxidation number / oxidation state :

→ It is a number assigned to an element according to AEN.

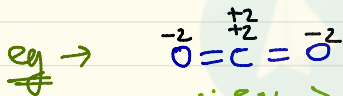
→ It is equal charge on an atom, if the bond is ionic.

→ If EN of A > EN of B

	O.S of A	O.S of B
A-B	-1	+1
A=B	-2	+2
A≡B	-3	+3



∵ EN_H < EN_{Cl} ∴ O.S of H = +1
O.S of Cl = -1



∵ EN_O > EN_C ∴ O.S of O = -2
O.S of C = +2 + 2 = +4

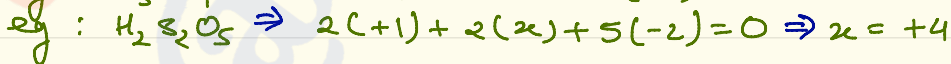
→ For ionic compounds oxidation state = Actual charge

eg: In NaCl, O.S of Na = +1, O.S of Cl = -1.

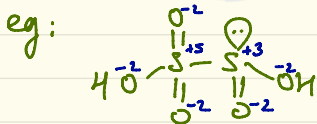
→ Average O.S = It calculated by using formula.

Sum of O.S of all atoms = net charge.

(O.S of H = +1) (O.S of O = -2)



→ Absolute O.S = It calculated by using structure.



(ix) Nature of oxides (A-O):

- Binary compounds of oxygen, in which oxidation state of oxygen is (-2).
- eg: Na_2O , K_2O , MgO , CO_2 , ZnO are oxides.
 F_2O is not an oxide, as O.S of 'O' = +2.
 K_2O_2 , H_2O_2 , BaO_2 are not oxides, as O.S of 'O' = -1
 NaO_2 , LiO_2 , KO_2 are not oxides, as O.S of 'O' = $-\frac{1}{2}$
- Metallic oxides, generally gives base in water so they are Basic oxides.
$$\text{Na}_2\text{O} + \text{H}_2\text{O} \longrightarrow \text{NaOH}$$

Basic Oxide Base
- Non Metallic oxide gives acids in water so they are acidic oxides.
$$\text{Cl}_2\text{O} + \text{H}_2\text{O} \longrightarrow \text{HOCl}$$

Acidic Oxide Acid
- Acidic nature of oxides \propto EN
 - \propto Oxidation state of A
 - \propto Non metallic nature of A
 - $\propto \frac{1}{\text{metallic nature of A}}$
- Left to Right, EN ↑ Basic nature ↓ Acidic nature ↑.
- Top to bottom,
 - for s-block, EN ↓ Basic Nature ↑ acidic nature ↓
 - for p-block, EN ↓ Basic Nature ↑ acidic nature ↓.
 - for d-block, EN ↑ Basic Nature ↓ acidic nature ↑.

eg:

Li_2O	BeO	B_2O_3	CO_2	N_2O_5	Basic Nature ↓ Acidic Nature ↑
Na_2O	MgO	Al_2O_3	SiO_2	P_4O_{10}	SO_2 Cl_2O_7
K_2O					Basic Nature ↓ Acidic Nature ↑
Rb_2O				Cu_2O	
				Ag_2O	Basic Nature ↓ Acidic Nature ↑
Cs_2O				Au_2O	
	Basic Nature ↑ Acidic Nature ↓				Basic Nature ↓ Acidic Nature ↑

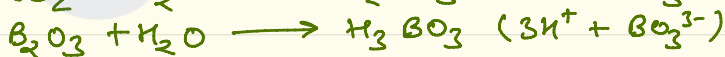
eg: Acidic nature of :

- (i) $\overset{+6}{\text{S}}\text{O}_3 > \overset{+4}{\text{S}}\text{O}_2$
 - (ii) $\overset{+5}{\text{N}}_2\text{O}_5 > \overset{+3}{\text{N}}_2\text{O}_3$
 - (iii) $\overset{+5}{\text{Sb}}_2\text{O}_5 > \overset{+3}{\text{Sb}}_2\text{O}_3$
 - (iv) $\overset{+4}{\text{Si}}\text{O}_2 < \overset{+5}{\text{P}}_2\text{O}_5 < \overset{+7}{\text{Cl}}_2\text{O}_7$
 - (v) $\text{SO}_3 > \text{SeO}_3 > \text{TeO}_3$
 - (vi) $\text{N}_2\text{O}_3 > \text{P}_2\text{O}_3 > \text{Sb}_2\text{O}_3$
- } O.S ↑ Acidic Nature ↑
- } non metallic character ↑
acidic nature ↑

Types of Oxides :

- (i) Acidic oxides: non metallic oxides are generally acidic oxides which reacts with base to give salt & water

Aqueous solution is acidic.



eg: SO_2 , P_2O_5 , N_2O_3 , N_2O_5 , CO_2 , SiO_2 , TeO_3 , SeO_3 etc

(ii) Basic Oxides: metallic oxides are generally basic oxides which react with acids to give salt & water
Aqueous solution is basic.



eg: Na_2O , K_2O , MgO , CaO , Cu_2O , Al_2O_3 etc.

(iii) Amphoteric Oxides: Oxides which react with acid as well as base to give salt and water

Aqueous solution is basic as well as acidic.



eg: जनाबे अली भारे मुना करे वो सभी असम के 3533 प्रभु

Zn	Be	Al	Ga	Sn	Cr ⁺³	V ⁺⁵	Sb ⁺³	As ⁺³		Pb
ZnO	BeO	Al ₂ O ₃	Ga ₂ O ₃	SnO SnO ₂	Cr ₂ O ₃	V ₂ O ₅	Sb ₂ O ₃	As ₂ O ₃		PbO PbO ₂

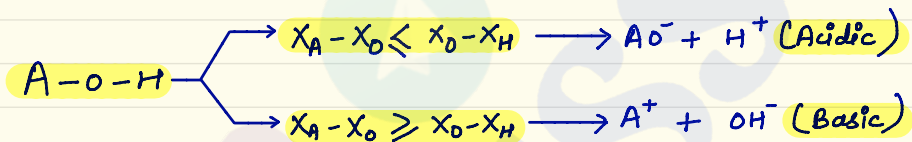
→ Metal oxides with (O.s = +4) ⇒ Amphoteric [VO_2 (Basic)]

(iv) Neutral Oxides: Oxides neither reacts with acid nor with base.

eg: CO , NO , NO_2

- Metallic oxides
 - +1, +2, +3 Basic except Amphoteric
 - +4 Amphoteric except VO_2
 - +5, +6, +7 Acidic except Amphoteric
- Non metallic oxides → Acidic except CO , NO , N_2O (neutral)
 H_2O (Amphoteric)
- Metalloids → Acidic except Amp. (As^{3+} , Sb^{3+})

(vii) Nature of Hydroxides (A-O-H):



EN of A ↓ $X_A - X_O$ ↑ basic nature of A-O-H ↑

eg: In NaOH
 $(X_O - X_{\text{Na}} = 2.6) > (X_O - X_H = 1.4)$
 therefore NaOH is basic in nature

eg: In ClOH
 $(X_O - X_{\text{Cl}} = 0.5) < (X_O - X_H = 1.4)$
 therefore ClOH is acidic in nature

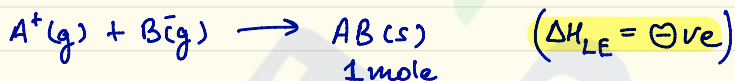
eg: Acidic nature of $\text{HOF} > \text{HOCl} > \text{HOBr} > \text{HOI}$
 down the group EN of halogen ↓ $X_O - X_A$ ↑ Acidic Nature ↓

exception: Acidic Nature of $\text{H}_3\text{PO}_4 < \text{H}_3\text{PO}_3 < \text{H}_3\text{PO}_2$
 (Based on Resonance)

Lattice Energy:

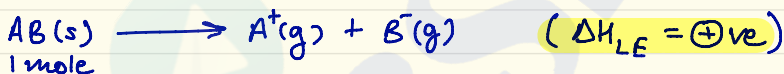
→ Definition 1:

Amount of energy released when one mole crystal lattice is formed from its constituent gaseous ions.



→ Definition 2:

Amount of energy required to break 1 mole of ionic lattice, into its constituent gaseous ions.



→ Mathematically:

$$E = \frac{k q_1 q_2}{r^2}$$

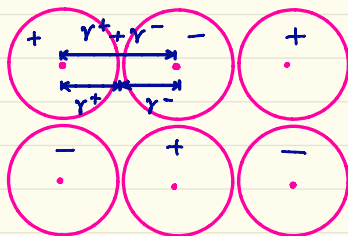
where, q_1 = charge on cation
 q_2 = charge on anion
 r = internuclear distance

→ Factors affecting Lattice energy (L.E):

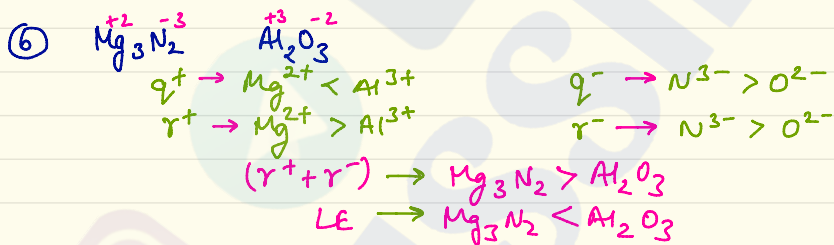
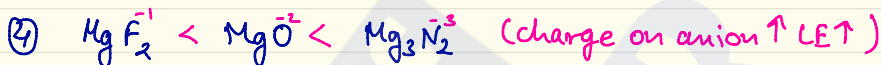
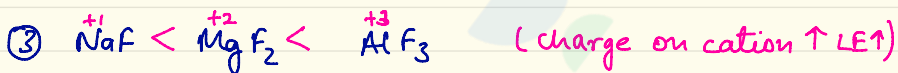
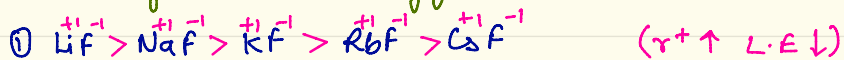
- 1.) Charge on ions \uparrow L.E \uparrow
 $LE \propto (q_1 \times q_2)$

- 2.) Size of ions \uparrow L.E \downarrow
 $LE \propto \frac{1}{(r^+ + r^-)}$

(priority of charge > size)

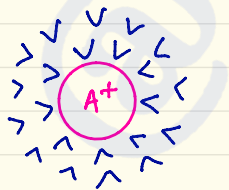
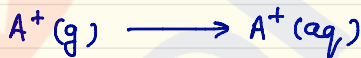


eg: order of Lattice energy :-

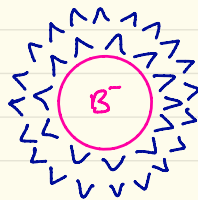
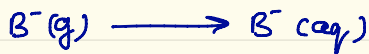


Hydration Energy:

Amount of energy released when one mole gaseous ion is completely dissolved in water.



ΔH_{HE} of A^+



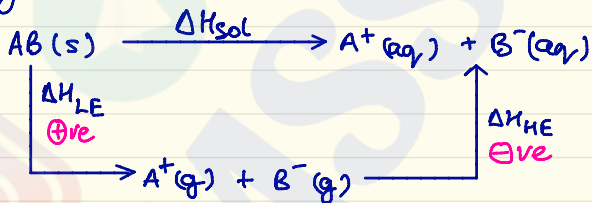
ΔH_{HE} of B^-

→ Factors affecting Hydration Energy:

- ① Hydration Energy \propto Charge of ions.
- ② Hydration Energy $\propto \frac{1}{\text{Size of ions}}$.

→ Applications of HE:

- ① Size of ion in aqueous medium \propto HE
- ② Mobility of Ions in aqueous medium $\propto \frac{1}{\text{HE}}$
- ③ Electrical Conductance $\propto \frac{1}{\text{Size of ion(aq)}} \propto \frac{1}{\text{HE}}$
- ④ Solubility \propto HE



→ An ionic compound is soluble if $\Delta H_{\text{HE}} > \Delta H_{\text{LE}}$.

eg: For alkali metal ions:

size of gaseous ions: $\text{Li}^+ < \text{Na}^+ < \text{K}^+ < \text{Rb}^+ < \text{Cs}^+$

Hydration Energy (ΔH_{HE}): $\text{Li}^+ > \text{Na}^+ > \text{K}^+ > \text{Rb}^+ > \text{Cs}^+$

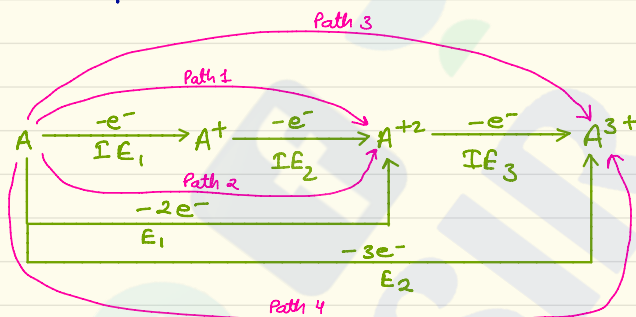
size of hydrated ions: $\text{Li}^+ > \text{Na}^+ > \text{K}^+ > \text{Rb}^+ > \text{Cs}^+$

Mobility of ions(aq): $\text{Li}^+ < \text{Na}^+ < \text{K}^+ < \text{Rb}^+ < \text{Cs}^+$

Electrical Conductance: $\text{Li}^+ < \text{Na}^+ < \text{K}^+ < \text{Rb}^+ < \text{Cs}^+$

Hess's Law: Energy is a state function and not path function. That is if there are more than one path for a conversion, then energy involved in each path is same.

eg:



from Hess's Law: Energy of path 1 = Energy of path 2

$$E_1 = IE_1 + IE_2$$

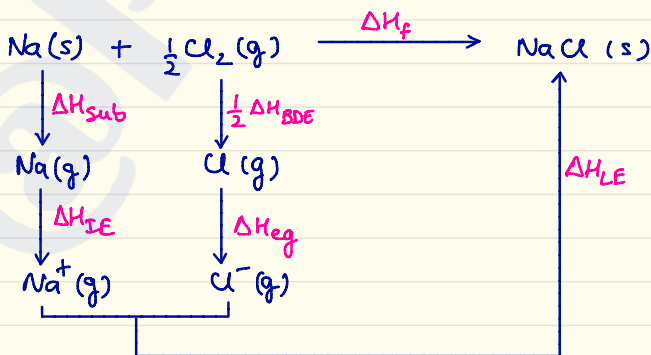
from Hess's Law: Energy of path 3 = Energy of path 4

$$E_2 = IE_1 + IE_2 + IE_3$$

Born Haber's Cycle:

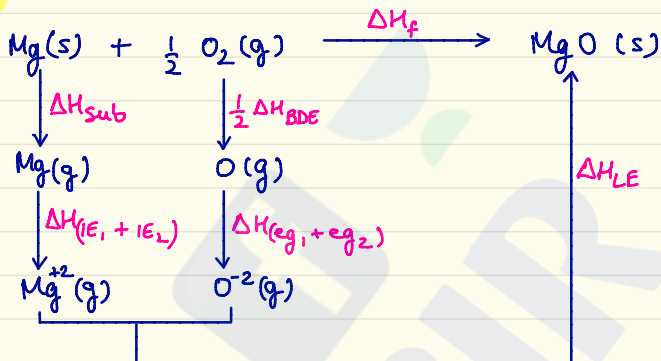
→ Born Haber's Cycle is used to calculate Lattice Energy.

→ for NaCl (s),



$$\Delta H_f = \Delta H_{\text{sub}} + \Delta H_{\text{IE}} + \frac{1}{2} \Delta H_{\text{BDE}} + \Delta H_{\text{eg}} + \Delta H_{\text{LE}}$$

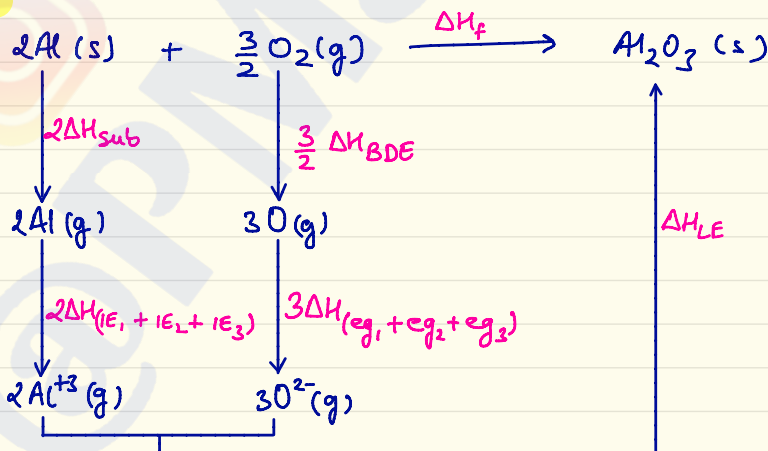
→ for MgO (s) ,



$$\Delta H_f = \Delta H_{\text{sub}} + \Delta H_{(\text{IE}_1 + \text{IE}_2)} + \frac{1}{2} \Delta H_{\text{BDE}}$$

$$+ \Delta H_{(\text{eg}_1 + \text{eg}_2)} + \Delta H_{\text{LE}}$$

→ for $\text{Al}_2\text{O}_3 (\text{s})$

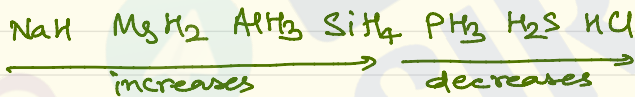


$$\Delta H_f = 2\Delta H_{\text{sub}} + 2\Delta H_{(\text{IE}_1 + \text{IE}_2 + \text{IE}_3)} + \frac{3}{2} \Delta H_{\text{BDE}}$$

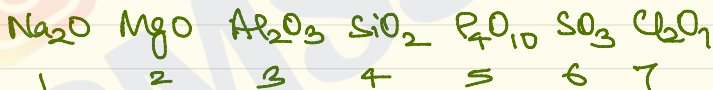
$$+ 3\Delta H_{(\text{eg}_1 + \text{eg}_2 + \text{eg}_3)} + \Delta H_{\text{LE}}$$

Valency: Combining capacity of atom

→ Valency with respect to Hydrogen



→ Valency with respect to oxygen



→ Based on O.S.

⁻⁷
ClH₇ not possible

⁺⁷
Cl₂O₇ possible