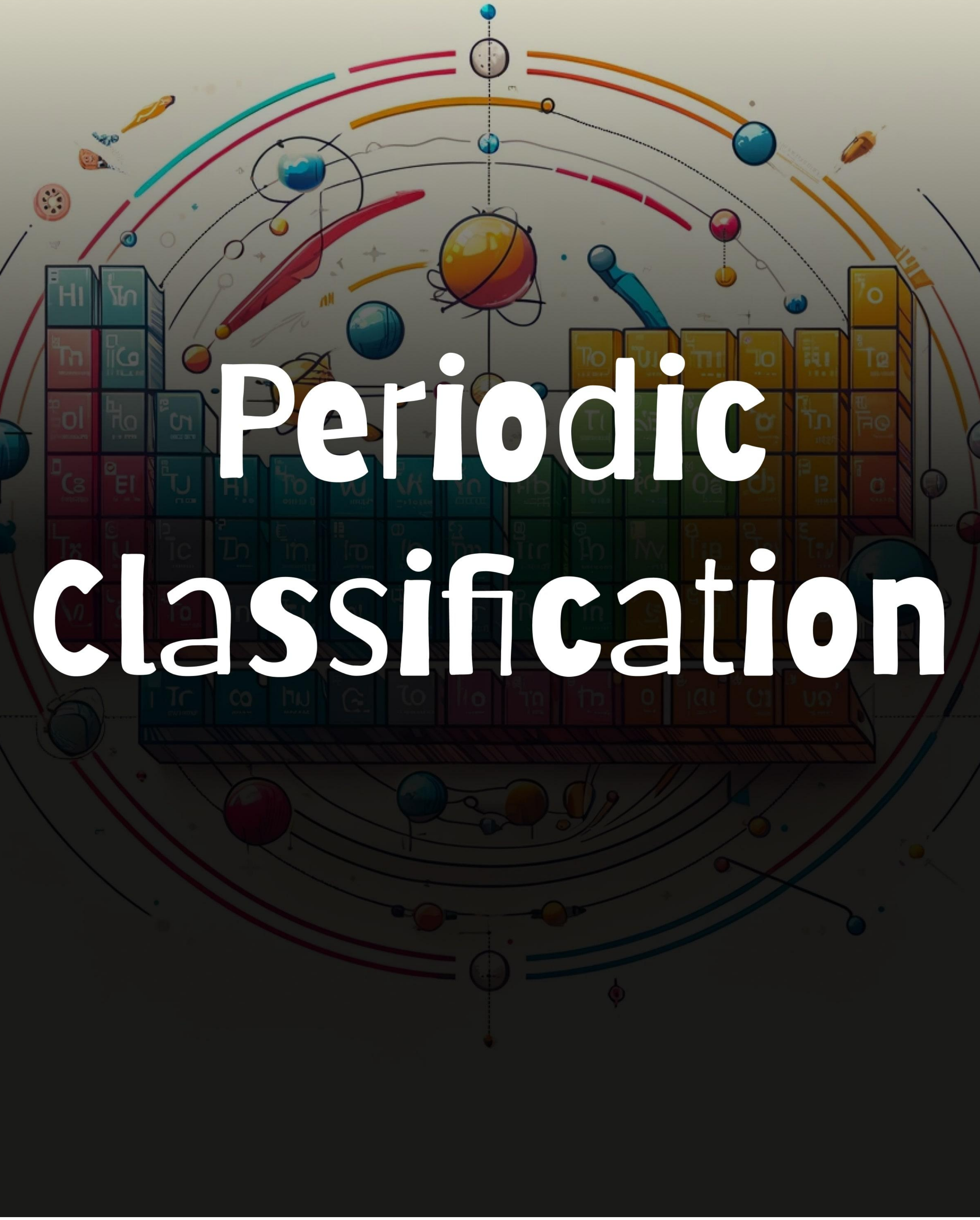


Periodic Classification

PeRiodic ClassiFication

PERIODIC CLASSIFICATION – Systematic arrangement of elements in tabular form is known as **periodic table.**

PROUT'S HYPOTHESIS

Atomic mass of elements is integer multiple of atomic mass of Hydrogen.
Atomic mass of any element = n·x mass of H.

E.g. H= n×1=4 - ---   Li= 7×1=7

But atomic mass of some elements was found to be fractional.

E.g. Cl= 35·5 ×1= 35·5 {not integer multiple}

After discovery of isotops, average mass of element come out to be fractional.
Even H itself has 3 isotops and its av. atomic mass is 1.008 which is fractional,
This theory was discarded.

DOBEREINER'S TRIADS

Dobereiner's arrange three elements (some properties) in increasing order of their atomic mass and found that atomic mass of middle element is arithmetic mean of terminal element.

1.

Li	Na	K
7	23	39

atom mass of Na = $\frac{39+7}{2}$ = 23

2.

Ca	Sr	Ba
40	88·5	137

3. Cl Br I

4. P As Sb

5. S Se Te

But C N O can not be Dobereiner's Triad.
12 14 16

as they don't have similar properties.

Dobereiner Triad is applicable only in groups (not in periods).

NEWLAND'S LAW OF OCTAWS

when elements are arranged in increasing order of their atomic masses, then properti of every eight element is similar to that of the first.

Li Be B C N O F

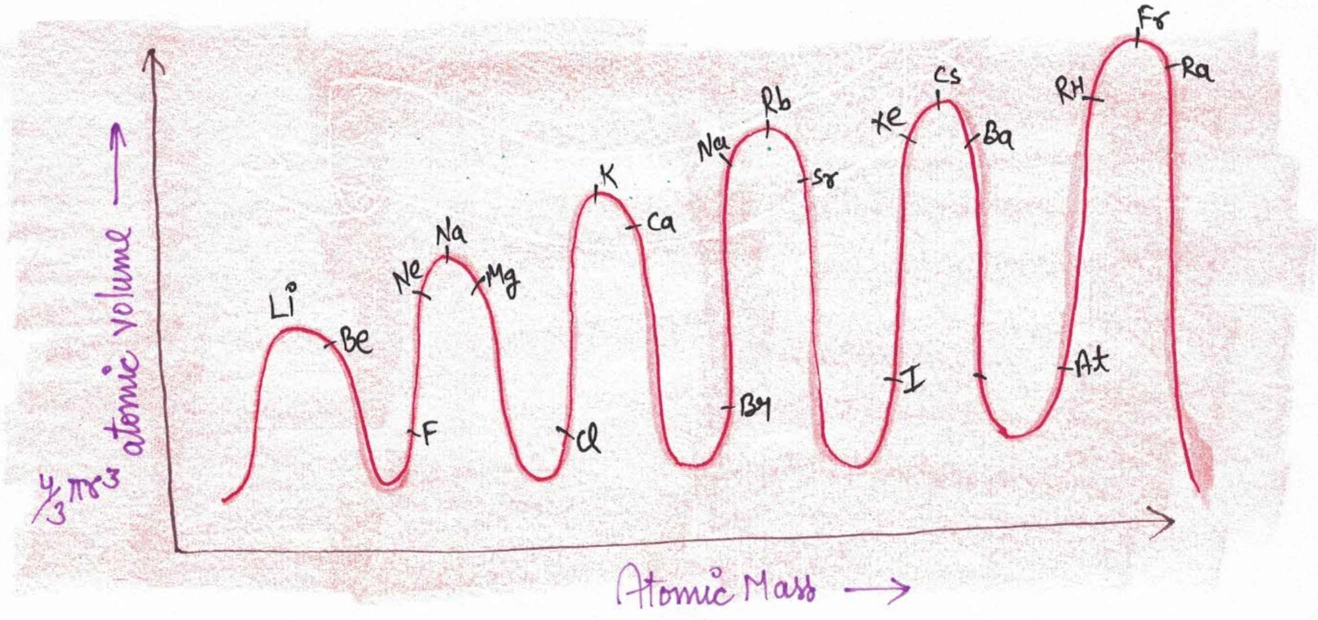
Na Mg Al Si P S Cl

K Ca

This law is applicable only upto Ca {atomic no, 20}.

LOTHER MEYER'S CURVE

Lother Meyer plotted a curve of atomic volume versus atomic mass.



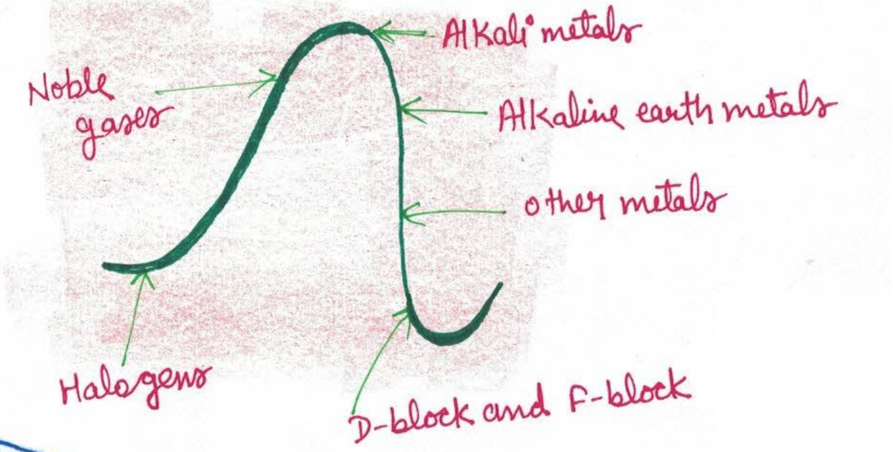
Lother Meyer found that same properties elements occupy same positions the curve.

Alkali metals [Li, Na, K, Rb, Cs, Fr] occupy peak position.

Alkaline earth metals [Be, Mg, Ca, Sr, Ba, Ra] occupy descending position on the curve.

Halogens (F, Cl, Br, I, Bt) and noble gases, [He, Ne, Ar, Kr, Xe, Rn] occupy ascending position.

D-block (transition elements) and F-block (inner transition elements) occupy bottom position.



MENDELEEV'S LAW

Physical and chemical properties of elements are periodic function of their atomic mass.

It means if elements are arranged in increasing order of their atomic mass periodicity occur after certain intervals.

Mendeleev's periodic table consist of two parts 7 horizontal rows and 9 verti-cle columns where 7 horizontal rows are periods and 9 verticles columns are groups.

1 Period,	2 elements	(H, He)	-	Very short period
2 Period,	8 elements	(Li to Ne)	-	short Period
3 Period,	8 elements	(Na to Ar)	-	short Period
4 Period,	18 elements	(K to Kr)	-	long Period
5 period,	18 elements	(Rb to Xe)	-	long Period
6 Period,	32 elements	(Cs to Rn)	-	Very Long period
7 Period,	Incomplete	(Fr to --)	-	Incomplete period

The group where divided into A & B subgroups

I A, B
II A, B
III A, B
IV A, B
V A, B
VI A, B
VII → Zero group

I		II		III		IV		V		VI		VII		VIII			0 (Zero)
A	B	A	B	A	B	A	B	A	B	A	B	A	B				
H																	He
Li	-	Be	-	B	-	C	-	N	-	O	-	F	-	-	-		Ne
Na	-	Mg	-	Al	-	Si	-	P	-	S	-	Cl	-	-	-		Ar
K	Cu	Ca	Zn	Ga	Sc	Ge	Ti	As	V	Se	Cr	Br	Mn	Fe	Co	Ni	Kr
Rb	Ag	Sr	Cd	In	Y	Sn	Zr	Sb	Nb	Te	Mo	I	Tc	Ru	Rh	Pd	Xe
Cs	Au	Ba	Hg	Tl	La	Pb	Hf	Bi	Ta	Po	W	At	Re	Os	Ir	Pt	Rn
Fr	-	Ra	-	-	Ac	-	-	-	-	-	-	-	-	-		-	-

MERITS OF MENDELEEV'S PERIODIC TABLE

Successful's classification

He corrected atomic mass of some elements Be, In, U

He left some places for elements which were not discovered at that time.

eka Boron → Scandium (Sc)
eka Aluminium → Gallium (Ga)
eka silicon → Germanium (Ge)
eka Hanganese → Technitium (Tc)

DEMERITS OF MENDELEEV'S PERIODIC TABLE

Position of Hydrogen is uncertain in Mendeleev's periodic table.
Hydrogen can be placed in :-

I A ↓ formscation (electropositive)
$H - 1e^- \rightarrow H^+$

IV A ↓ shars e^-
$H - H \rightarrow H_2$

VII A ↓ electronegative native

Position of isotops uncertain in mendeleev's Periodic Table. Isotops were placed at same position in periodic table.

$_1H^1$, $_1H^2$, $_1H^3 \longrightarrow$ I A

$_6C^{12}$ $_6C^{13}$ $_6C^{14} \longrightarrow$ IV A (2nd period)

isotops
↙ Z same
↘ A (mass) different

Position of isobars $_{18}Ar$ and $_{20}Ca$ should be placed in same place but they are placed at diff. position.

$_{18}Ar^{40}$, $_{20}Ca^{40}$

Same properties of elements were placed in different group. Ba and Pb shows similar properties.

Ba (II A) and Pb (IV A)

Different properties of elements were placed in same group.

E.g
I.A → Alkali metals (s-block)
I.B → Coinage metals (d-block)

They have very diff. properties except same oxidation state.

position of lanthanides and actinides
↘ bottom of the table.

Anomalous Pair

$_{18}Ar$	$_{19}K$	$_{52}Te$	$_{53}I$
39.9	39	12.7	$12^{6}.9$
$_{27}Co$	$_{28}Ni$	$_{90}Th$	$_{91}Pa$
58	59.5	231	230

4 Places where mendeleev's Law is violated. gave a lot of on volency factor.

MODERN PERIODIC TABLE

MODERN PERIODIC LAW ⇒ physical and chemical properties of elements are period function of their atomic number (no. of protons. as no. of e-).

Modern periodic table is based on electronic configuration.

Modern periodic table consist of two parts :-

1. 7 horizontal rows (known as periods)
2. 18 vertical column (known as groups)

PERIODS

1st Period	$_{1}H^{1}$ -	$_{2}He$	(1s)	2 element
2nd Period	$_{3}Li$	$_{10}Ne$	(2s+2p)	8 element
3rd Period	$_{11}Na$	$_{18}As$	(3s+3p)	8 element
4th Period	$_{19}K$	$_{36}Kr$	(4s+3d+4p)	18 element
5th Period	$_{37}Rb$	$_{54}Xe$	(5s+4d+5p)	18 element
6th Period	$_{55}Cs$	$_{86}Rn$	(6s+4f+5d+6p)	32 element
7th Period	$_{87}Fr$	$_{118}Uuo$	(7s+5f+6d+7p)	32 element

GROUPS

S block – 1 and 2
P block – 13 to 18
d block – 3 to 12
f block – 3rd group

Modern Periodic table divided into 4 blocks :-

S - block

Last e^- or differentiating e^- enters into s - subshell.
I and II groups elements are known as s block elements.
s - block elements are known as representative elements, they represent left side of the periodic table. Also known as normal elements
As last e^- enters into valence shell.
IA group elements (except H) known as alkali metals. as they dissolve on in cold water to form soluble hydroxide (alkali) and give H_2 gas.

$$2M + 2H_2O \longrightarrow MOH + H_2$$

{M → Li, Na, K, Rb, Cs} {Fr is radioactive}

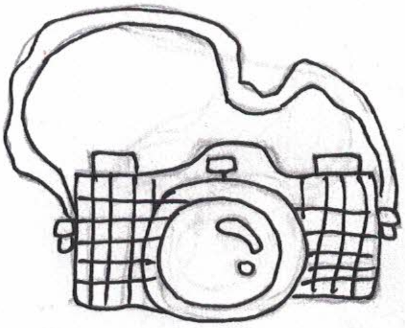
MECHANISM - $2Li^0 \longrightarrow 2e^- + 2Li^+$ (oxidation)

$$2H_2O \longrightarrow 2H^+ + 2OH^-$$

$$2H^+ \longrightarrow 2OH^- + LiOH$$

$$2H^+ \longrightarrow 2e^- + H_2 \quad \text{(Reduction)}$$

$$\Rightarrow 2Li^0 + 2H_2O \longrightarrow 2LiOH + H_2 \quad \text{(Redox)}$$



IA group elements are → Li, H, Na, K, Rb, Cs, Fr
H → non - metal, exist in gas p
Li, Na, K, Rb - exist in solid state
C, Fr are liquid at room temperature
Fr is radio element.

Electronic configuration of I group — ns
n → no. of shell

IA group elements shows (+1) oxidation state in their compound.
IA group elements are paramagnetic in nature but their ions are diamagnetic in nature.

Paramagnetic — unpaired e^- is present
Diamagnetic — paired e^- is present

E.g. Na → $1s^2, 2s^2, 2p^6, 3s^1$ → Paramagnetic
Na^+ → $1s^2, 2s^2, 2p^6$ → Diamagnetic

Reactivity of I group elements increase down the group.

order of Reactivity → Li Na K < Rb < Cs

Reactivity ∝ electropositivity

Alkali metals imparts colour to bunsen frame when alkali metals are brought near flame then they absorb energy and transit to excited state but at higher energy level they are unstable so they come back to then ground state and in this process they emit energy which is form of colour.

Li → crimson Red

Na → Golden yellow

K → lilac

Rb → Reddish violet

Cs → violet

II A group element are known as alkaline earth metals. They are found in earth's crust in the form of oxides. Their oxides dissolve in water to form hydroxides

$$CaO + H_2O \longrightarrow Ca(OH)_2$$

$$MO + H_2O \longrightarrow M(OH)$$

Electronic configuration of IIA group elements is ns^2.

Elements of IIA group are - Be, Mg, Ca, Sr, Ba, Ra.

Radium is radioactive element.

All elements exist in solid state.

Alkaline earth metals and their ions are diamagnetic in nature.

$$ns^2 - 2e^- \longrightarrow ns^0$$

E.g - $Mg - 2e^- \longrightarrow Mg^{+2}$

$1s^2 2s^2 2p^6 3s^2$ $\quad$ $1s^2 2s^2 2p^6$

Reactivity of alkaline earth metals increases down the group.

$$Be < Mg < Ca < Sr < Ba$$

All metals except Be and Mg imparts colour to bunsen flame.

Ca → brick red

Sr → Crimson red

Ba → apple green

P - Block

- position - right side of a periodic table.
- Also known as normal elements as last e^- enters into
- outermost shell is $ns^2\ np^{1-6}$ {n → no. of shell}
- P- block elements exist in gas phase liquid phase and solid phase.
- P- block elements are metals, non-metals and metalloids.
- P- block elements shows paramagnetic and diamagnetic nature.

13. (III A) Baron family ($ns^2 np^1$).

(B) Baron – metalloid – solid
(Al) Alluminium – metal – solid
(Ga) Gallium – metal – Liquid
(In) Induium – metal – solid
(Th) Thalium – metal – solid

14 Group (IVA) Carban family ($ns^2 np^2$)

Carbon (C) – non-metal – O.S → +4 to −4
Silicon (Si) – metalloid – O.S → +4
Germanium (Ge) – metalloid – O.S → +2, +4
Tin (Sn) – non-metal – O.S → +2, +4
Lead (Pb) – non-metal – O.S → +2, +4

15th Group (VA) Nitrogen family (ns^2, np^3)

Nitrogen (N) – non-metal (gas) → −3 to +5
Phosphorus (P) – non-metal (solid) →
Arsenic (As) – metalloid (solid) →
Stibuim (Sb) – metalloid (solid) →
Bismuth (Bi) – metal (solid) →
+3 to +5

Nitrogen exist in diatomic form and others as tetra atomic form (P_4, As_4)
chalcogens (ore forming elements)

16th Group (VIA) Oxygen family (ns^2, np^4)

Oxygen (O) –
Sulphur (S) –
Selenium (Se) –
non metal
Tellinium (Te) –
metalloids
Polonium (Po) –
Radioactive elements

O.S. of oxygen → −2 to +2

K_2O	K_2O_2	KO_2	O_2	O_2F_2	OF_2
(−2)	(−1)	(−½)	(0)	(+1)	(+2)

O.S of sulphur and selezium → −2 to +6

Te and Po → +2, +4

Oxygen exist in gas phase (O_2).

17th group (VII A) - Halogens (salt forming)

electronic configuration (ns^2, np^5)

Flourine (F)	non metals	gas
Chlorine (Cl)		gas
Bronine (Br)		liquid
Iodine (I)		solid
Astatine (At)	metalloid	solid - radioactive

colour

F_2 - Yellow gas

Cl_2 Yellow green gas

Br_2 Red orange gas

I_2 Violet

Reactivity down the group decreases

$F > Cl > Br > I$

Reason - (oxidising nature down the group decreases)

F is the strongest oxidising agent in the periodic table.

18th group (VIII A) - Inert gas (ns^2, p^6)

He - $1s^2$

Ne - [He] $2s^2 3p^6$

Ar - [Ne] $3s^2 3p^6$

Kr - [Ar] $3d^{10} 4s^2 4p^6$

Xe - [Kr] $4d^{10} 5s^2 5p^6$

Rn - [Xe] $4f^{14} 5d^{10} 6s^2 6p^6$

Uuo - [Rn] $5f^{14} 6d^{10} 7s^2 7p^6$

D - Block

- position - 3 to 12 group
- Last e^- enters into penultimate d subshell
- electronic configuration of d-block is

[Inert gas] $(n-1)d^{1-10} ns^{1-2}$

- $(n-1)$ → penultimate shell

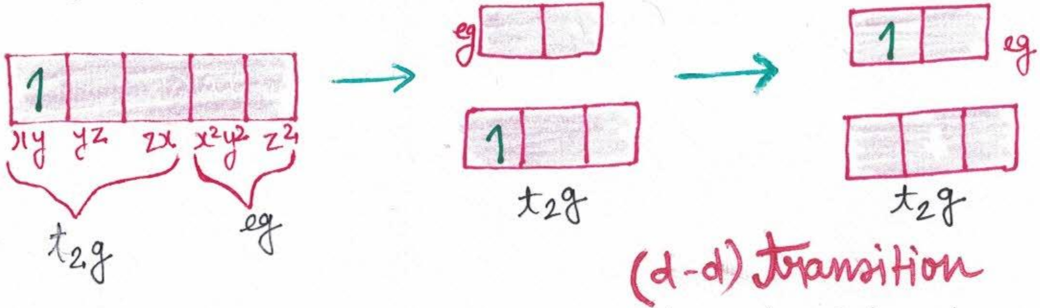
(n) → outermost shell

- Four d-block series

I	3d →	$_{21}Sc$ to $_{30}Zn$	$3d^{1-10} 4s^{1-2}$
II	4d →	$_{21}Y$ to $_{48}Cd$	$4d^{1-10} 5s^{1-2}$
III	5d →	La, $_{72}Hf$ to $_{80}Hg$	$5d^{1-10} 6s^{1-2}$
IV	6d →	$_{89}Ac$, $_{104}Rf$ to $_{112}Uub$	$6d^{1-10} 7s^{1-2}$

All d block elements are metal and all exist in solid state except mercury (liquid).

d-block elements forms coloured compounds because unpaired e^- present in d-subshell. when unpaired e^- present in d-subshell.
(d-d) transition takes place.



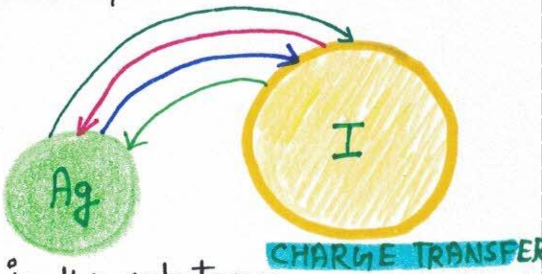
(d-d) transition

iff all e^- are paired in d-subshell, charge transfer takes place.

AgI (silver iodide) ⟶ yellow colour

$_{47}Ag \longrightarrow {}_{36}[Kr]\ 5s^1\ 4d^{10}$

$Ag^+ \rightarrow {}_{36}[Kr]\ 5s^0\ 4d^{10}$



7. d-block elements are used as catalyst. catalyst is the substance that alters the rate of reaction which increases the rate of reaⁿ → (+) catalyst which decrease the rate of reaⁿ → (–) catalyst

(i) d block elements show variable vallency
have same size
they have vacant orbitals in s, d-subshell

8. d-block elements shows variable oxidation state $(n-1)d^{1-10}ns^2$ [Penultimate d-subshell or outermost e^- takes part in recⁿ]

E.g. $Z = 25\ (Mn) \rightarrow {}_{18}[Ar]\ 3d^5\ 4s^2$

$Mn^{+2} \rightarrow 3d^5$ $Mn^{+3} \rightarrow 3d^4$ $Mn^{+4} \rightarrow 3d^3$
$Mn^{+5} \rightarrow 3d^2$ $Mn^{+6} \rightarrow 3d^1$ $Mn^{+7} \rightarrow 3d^0$

$MnO \rightarrow (+2)$ $MnO_2 \rightarrow (+3)$ $Mn_2O_3 \rightarrow (+3)$
$MnO_2 \rightarrow (+4)$ $Mn_2O_5 \rightarrow (+5)$ $MnO_3 \rightarrow (+6)$ $Mn_2O_7 \rightarrow (+7)$

Maganese shows maximum no. of oxidation state
osmium (Os) and Ruthenium (Ru) shows highest oxidation state (+8).

E.g. osmium tetraoxide OsO_4

9. d-block elements forms alloys as size is almost same.

ALLOY = mixture of two or more than 2 metals

Brass → Cu + Zn
Bronze → Cu + Sn
Steel → Fe + Cr + Ni + Mn

d-block elements and their ions are generally paramagnetic in nature.

Fe^{+3} = 18 [Ar] $3d^5 . 4s^0$ $\quad$ Fe^{+2} = 18 [Ar] $3d^6 4s^0$

Cu^{+2} = 18 [Ar] $3d^9 4s^0$ $\quad$ Zn^{+2} = 18 [Ar] $3d^{10} 4s^0$

diamagnetic

f - block

- position → III B or 3rd group.
  IV and VII period

Known as inner transition elements because last e- enters into (n-2) f subshell or pre- penultimate subshell.

electronic configuration

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

f - block → Lanthanide (4f series) $_{58}Ce$ to $_{71}Lu$
f - block → Actinide (5f series) $_{90}Th$ to $_{103}Lr$

All f-block elements are metals and exist in solid state.

General oxidation state of f-block elements is (+3).

All actinides are radioactive element. only Pm (Promethium) is radioactive in Lanthanides.

Determination of Group Period and Block

S-block

Electronic configuration of element is [inert] ns^1 means 1st group or [inert] ns^2 means 2nd group.

Period → highest principal quantum number.

E.g. Z = 38 = $_{36}[Kr]\, 5s^2$
Period → 5 $\quad$ group → 2

Z = 87 = $_{86}[Rn]\, 7s^1$
Period → 7 $\quad$ group → 1

P-block

Electronic configuration → $ns^2 + (n-1)d^{10} + np^{1-6}$
group number = 12 + np e^{-}s
Period = highest "n"

E.g. Z = 32 = $_{18}[Ar]\, 4s^2\, 3d^{10}\, 4p^2$
group no. → 12+2 = 14 $\quad$ period → 4

Z = 81 = $_{54}[Xe]\, 6s^2\, 4f^{14}\, 5d^{10}\, 6p^1$
group no. → 12+1 = 13 $\quad$ period = 6

Z = 53 = $_{36}[Kr]\, 5s^2\, 4d^{10}\, 5p^5$
group no. → 12+5 = 17 $\quad$ period → 5

d-block

Electronic configuration is [inert] $ns^{1-2}(n-1)d^{1-10}$

group no. → no. of n-s, e^- + $(n-1)$ d e^{-s}

period → highest "n"

E.g. 1. Z = 24 = $_{18}[Ar]\,4s^1\,3d^5$

group no → 1+5 = 6

Period → 4

2. Z = 78 = $_{54}[Xe]\,6s^1\,4f^{14}\,5d^9$ platinum

group no = 1+9 = 10

Period = 6

3. Z = 48 = $_{36}[Kr]\,5s^2\,4d^{10}$

group → 2+10 = 12 (IIB)

Period → 5

4. Z = 42 = $_{36}[Kr]\,5s^1\,4d^5$

group no. = 5+1 = 6

Period = 5

5. Z = 29 = $_{18}[Ar]\,4s^1\,3d^{10}$

group no. = 11 (IB)

period = 4

f-block

Electronic configuration → [inert] $(n-2)f^{1-14}\,(n-1)\,d^{1-10}\,ns^2$

group no = 3

period no. VI [58 to 71]

VII [90 to 103]

Q. Z = 43 determine Z above 43 and below 43.

Z above 43 ⇒ 43 − 18 = 25

Z below 43 ⇒ 43 + 32 = 75

as Z = 43, is a d-block element.

25 Mn
43 Tc
75 Re

Q. Z = 48

Z above ⇒ 48 − 18 = 30

Z below ⇒ 48 + 32 = 80

Zn
Cd
Hg

Q. Z = 33

Z above ⇒ 33 − 18 = 15

Z below ⇒ 33 + 18 = 51

Z below 51 ⇒ 51 + 32 = 83

P
As
Sb

Q. Z = 49

Z above ⇒ 49 − 18 = 31

Z below ⇒ 49 + 32 = 81

Ga
In
Tl

Q. $Z = 38$

Z above ⇒ $39 - 18 = 21$ Sc

Z below ⇒ $39 + 18 = 57$ La

Q $Z = 38$

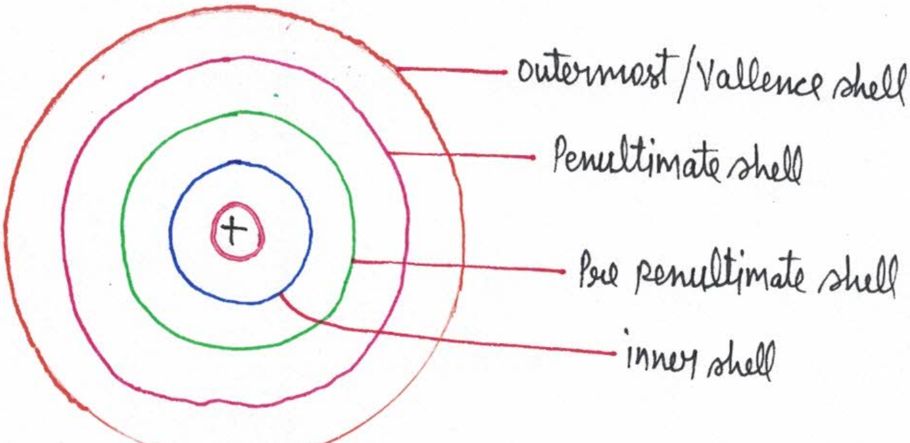
Z above ⇒ $38 - 18 = 20$

Z below ⇒ $38 + 18 = 56$

s	d	p
2		
8		8
8	18	18
18	18	18
18	32	32
32	32	32

EFFECTIVE NUCLEAR CHARGE ZEFF

Z = charge = no. of protons



In multielectron system, all inner shell electrons repeat outermost electrons, therfor nuclear force of attraction towards outermost electron decrease. Resultant force of attraction experienced by outermost electron is known as effective nuclear charge.

$$Z_{effective} = Z - \sigma$$

Z = atomic number

σ = screening constant

σ outermost → 0.35

σ Penultimate → 0.85

σ pre penultimate → 1

σ inner shell → 1

Slater's Rule for assigning screening constant (σ).

1. σ for 1s e⁻ of H and He is 0.30.
2. σ for all ns and np (outermost) shell e⁻ = 0.35
3. σ for penultimate shell e⁻s = 0.85
4. σ for all inner e⁻s = 1.0

5. σ for nd and nf electron (if they are vallence electron) = 0·35

All remaining $e^- = \sigma = 1{\cdot}0$

σ = (no. of vallence shell electron − 1) × 0·35
+ all penultimate shell e^-s) × 0·85
+ (Remaining e^-) × 1

Ist Period

$H = 1s^1$ $\sigma = (1-1) \times 0{\cdot}30 = 0$

Screening effect or penultimate effect is not applicable on H-atom.

$Z_{eff} = Z - b = 1 - 0 = 1$

$He = 1s^2$ $\sigma = (2-1) \times 0{\cdot}30 = 0{\cdot}30$

$Z_{eff} = Z - \sigma = 2 - 0{\cdot}30 = 1{\cdot}7$

when we move left to right in period, sigma decrases and zeff increases. Zeff is dominating over screening constant.

IInd Period

$_3Li = 1s^2\, 2s^1$ $\sigma = (1-1) \times 0{\cdot}35 + (2) \times 0{\cdot}85$

$\sigma = 1{\cdot}7$

$Z_{eff} = Z - \sigma = 3 - 1{\cdot}7 = 1{\cdot}3$

$_4Be = 1s^2\, 2s^2$ $\sigma = (2-1) \times 0{\cdot}35 + 2 \times 0{\cdot}85$

$\sigma = 0{\cdot}35 + 1{\cdot}7 = 2{\cdot}05$

$Z_{eff} = Z - \sigma = 4 - 2{\cdot}05 = 1{\cdot}95$

$_5B = 1s^2\, \underbrace{2s^2\, 2p^1}$ $\sigma = (3-1)0{\cdot}35 + (2)0{\cdot}85$

$\sigma = (2)\, 0{\cdot}35 + 2(0{\cdot}85) = 0{\cdot}7 + 1{\cdot}7 = 2{\cdot}4$

$Z_{eff} = 5 - 2{\cdot}4 = 2{\cdot}6$

$_6C = 1s^2\, 2s^2\, 2p^2$ $\sigma = (4-1)\, 0{\cdot}35 + (2)\, 0{\cdot}85$

$\sigma = 1{\cdot}75 + 1{\cdot}7 = 3{\cdot}45$

$Z_{eff} = 6 - 2{\cdot}75 = 3{\cdot}25$

$_7N = 1s^2\, 2s^2\, 2p^3$ $\sigma = (5-1)\, 0{\cdot}35 + 2 \times 0{\cdot}85$

$\sigma = 1{\cdot}40 + 1{\cdot}7 = 3{\cdot}1$

$Z_{eff} = 7 - 3{\cdot}1 = 3{\cdot}9$

$_8O = 1s^2\, 2s^2\, 2p^4$ $\sigma = (6-1)\, 0{\cdot}35 + 2 \times 0{\cdot}85$

$\sigma = 1{\cdot}75 + 1{\cdot}7 = 3{\cdot}45$

$Z_{eff} = 8 - 3{\cdot}45 = 4{\cdot}55$

$_9F = 1s^2 2s^2 2p^5$ $\sigma = 6 \times 0.35 + 2 \times 0.85$

$\sigma = 2.10 + 1.7 = 3.8$

$Z_{eff} = 9 - 3.8 = 5.2$

$_{10}Ne = 1s^2 2s^2 2p^6$ $\sigma = 7 \times 0.35 + 1.7$

$\sigma = 2.45 + 1.7 = 4.15$

$Z_{eff} = 10 - 4.15 = 5.85$

GROUPS

$_{11}Na = 1s^2 2s^2 2p^6 3s^1$ $\sigma = (1-1)0.35 + (8)0.85 + (2)1$

$\sigma = 6.80 + 2 = 8.80$

$Z_{eff} = 11 - 7.8 = 2.2$

$12 = 1s^2 2s^2 2p^6 3s^2$ $\sigma = 0.35 + 8 \times 0.85 + 2$

$\sigma = 0.35 + 2 - 6.8 = 9.15$

$Z_{eff} = 12 - 9.15 = 2.85$

Left to Right across the period. σ increases by 0.35 per e^- and Zeff increases by 0.65 per e^-

Zeff > screening effect

Group

$_{19}K = 1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$ $\sigma = 0 + 8(0.85) + 10$

$\sigma = 6.8 + 10 = 16.8$

$Z_{eff} = 19 - 16.8 = 2.2$

Trends in Groups

	σ	Zeff
H	0	1
Li	1.7	1.3
Na	8.8	2.2
K	16.8	2.2
Rb	34.8	2.2
Cs	52.8	2.2
Fr	84.8	2.2

Down the group, Zeff first increases and then remains constant. Sigma gradually increases but Zeff first increases and then remains constant.

In group σ is dominating factor.

Q. $_{29}Cu =$ $\boxed{\text{a}}$ σ for 4s e⁻ $\boxed{\text{b}}$ σ for 3d e⁻

$\boxed{\text{c.}}$ Zeff for 4s e⁻ $\boxed{\text{d.}}$ Zeff for 3d e⁻

S. $_{29}Cu = {}_{18}[Ar]\,4s^1 3d^{10} = 1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^1$

$\boxed{\text{a.}}$ $\sigma = (1-1) \times 0{\cdot}35 + (18)\,0{\cdot}85 + (10)$

$= 0 + 15{\cdot}3 + 10 = 25{\cdot}3$

$\boxed{\text{c.}}$ Zeff $= 29 - 25{\cdot}3 = 3{\cdot}7$

$\boxed{\text{b.}}$ L[illegible] e⁻ enters into 3d

$_{29}Cu = 1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10}$

$\sigma = [(\text{no. of } nd + nf \text{ e}^-\text{s}) - 1]\,0{\cdot}35 + (\text{Remaining e}^-) \times 1$

$\sigma = (10-1)\,0{\cdot}35 + 18 \times 1$

$= 9 \times 0{\cdot}35 + 18 = 3{\cdot}15 + 18 = 21{\cdot}15$

$\boxed{\text{d.}}$ Zeff $= 29 - 21{\cdot}15 = 7{\cdot}85$

∴ 3d e⁻ are tightly held by nucleus than 4s e⁻
we'll remove e⁻ from 4s to form cation.

Q. $_{58}Ce = {}_{54}[Xe]\,6s^2 4f^1$ $\cdot\, 4f^2$

$= 1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 4d^{10} \underbrace{5s^2 5p^6} 6s^2$

S. σ for S e⁻ $= (2-1)\,0{\cdot}35 + (8)\,0{\cdot}85 + (48)\,1.$

$= 0{\cdot}35 + 6{\cdot}80 + 48$

$= 55{\cdot}15$

Zeff $= 58 - 55{\cdot}15 = 2{\cdot}85$

σ for 4f $= 1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 4d^{10} 4f^2$

$= (12-1)\,0{\cdot}35 + 36$

$= 11 \times 0{\cdot}35 + 36 = 3{\cdot}85 + 36 = 39{\cdot}85$

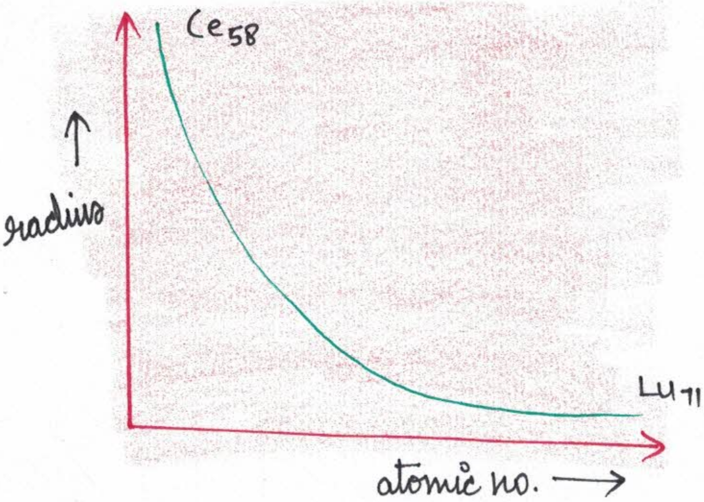
Zeff $= 58 - 39{\cdot}85 = 18{\cdot}15$

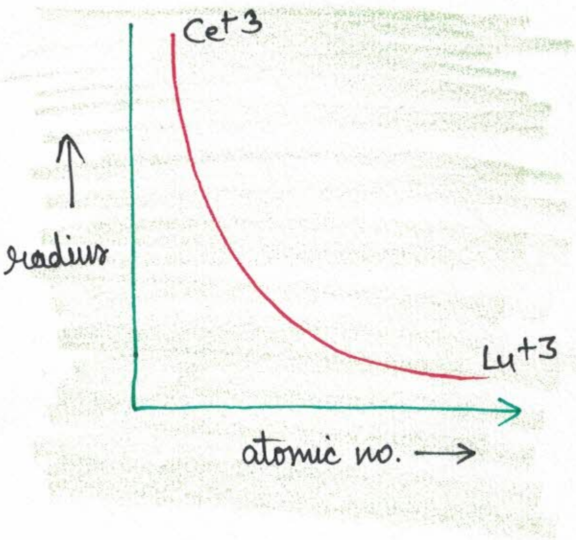
∴ pre penultimate shell (4f) is more tightly held by nucleus.

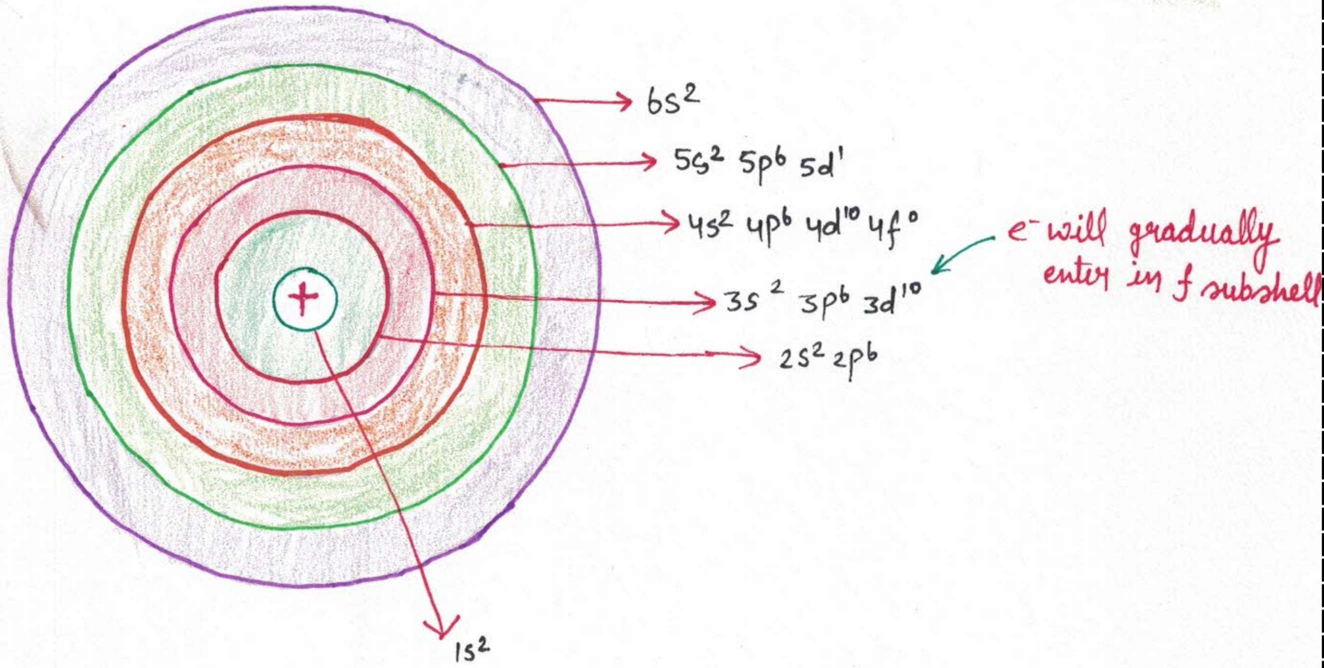
Zeff of 4f > Zeff of 4d > Zeff of s and p

LANTHANIDE CONTRACTION

In Lanthanide series, electrons enters into 4f subshell 4f subshells are diffused in shape. electronic repulsion b/w 4f subshell e^- is 4f subshell is closer to nucleus. Therefore, 4f electrons feel more force of attraction by the nucleus. size of lanthanide decrease from $_{58}Cr$ to $_{71}Lu$. This is phenominon from $_{58}Cr$ to $_{71}Lu$. This is phenomenon is known as Lanthanide contraction.







Consequences of Lanthanide Construction

① Atomic and ionic radii of lanthanides. the atomic and ionic radii of lanthanides decreases due to lanthanide contraction.

$$\text{Lanthanide contraction} \rightarrow \text{radii decrease}$$
$$\searrow \text{Zeff increases}$$

② Acidic nature of lanthanum hydroxide

$Ln(OH)_3$ $\quad$ $Ln \rightarrow$ Ce to Lu

$Ce(OH)_3$, $Pr(OH)_3$ $\quad$ $Nd(OH)_3$ $\quad$ $Sm(OH)_3$ ---- $Lu(OH)$

Ce^{+3} stable

Basic in nature

Lu^{+3} unstable (It easily gains electrons) e^- accept

acidic nature

Cerium is more basic in nature and Lutetium is acidic in nature.
All Lanthanides are metals and metallic hydroxides are basic in nature. But cerium hydroxides to Lutetium hydroxide. acidic nature of hydroxides increases due to Lanthanide contraction.

$Ce(OH)_3$ -- < $Pr(OH)_3$ < $Nd(OH)_3$ < --- < $Lu(OH)_3$

basic $\qquad\qquad\qquad\qquad$ most acidic

③ size of 4d series element is almost equal to size of 5d series elements. due to lanthanide contraction.

3d	Sc	Ti	Y	Cr	Mn	Fe	Co	Ni	Cu	Zn
4d	$\hat{Y}$	$\hat{Zr}$	$\hat{Nb}$	$\hat{Mo}$	$\hat{Tc}$	$\hat{Ru}$	$\hat{Rh}$	$\hat{Pd}$	$\hat{Ag}$	$\hat{Cd}$
5d	$\overset{\simeq}{La}$	$\overset{\simeq}{Hf}$	$\overset{\simeq}{Ta}$	$\overset{\simeq}{W}$	$\overset{\simeq}{Re}$	$\overset{\simeq}{Os}$	$\overset{\simeq}{Ir}$	$\overset{\simeq}{Pt}$	$\overset{\simeq}{Au}$	$\overset{\simeq}{Hg}$
6d	Ac									

size of Hf $\simeq$ size of Nb

$$3d < 4d \simeq 5d$$

size ↑ due to no. of shells will increase from 4d to 5d but due to lanthanide contraction size will decrease.

so, these two factors will counter balance each other and size of 4d is almost similar to 5d element.

④ Ionisation potential, Amount of energy required to remove one e^- from outermost shell of neutral isolated gaseous atom, is known as Ionisation Potential.

3d > 4d > 5d   I P order

PERIODICITY

ATOMIC RADIUS

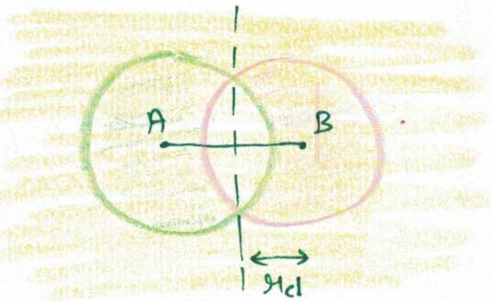
Distance b/w the nucleus and outermost shell of an atom is known as atomic radius.

unit of atomic radius → unit A°. $1A° = 10^{-10} m$

or piometer $1PM = 10^{-12} m$

- Covalent Radius
- Ionic Radius
- Vander wall's Radius
- Metallic Radius

1. COVALENT RADIUS



Half of the internuclear distance b/w two covalent bonded atoms in a molecule is known as covalent radius.

A. For homoatomic molecules $= r_c = \frac{d_{AB}}{2}$

E.g. $H_2, Cl_2, O_2, N_2, Br_2$

$d_a = 15.4A°$

$r_c = \frac{1.54}{2} = 0.77A°$

E.g. $d_{Cl-Cl} = 1.98A°$

$r_{Cl} = \frac{1.98}{2} = 0.99A°$

B. For hetroatomic molecule

$$d_{AB} = r_A + r_B - 0.09(X_A - X_B)$$

$d_{C-C} = 1.54 A°$ $d_{Cl-Cl} = 1.98 A°$

$X_C = 2.5$ $X_{Cl} = 3$

$r_C = 0.77 A°$ $r_{Cl} = 0.99 A°$

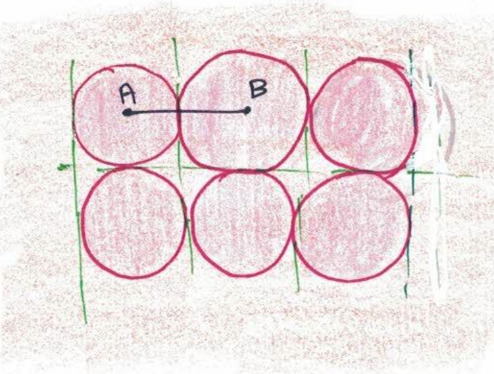
$$d_{AB} = 0.77 + 0.99 - 0.09(+2.5-3)$$

$$d_{Cl-Cl} = 1.76 - 0.09 \times 0.5$$

$$= 1.760 - 0.045$$

$$= 1.715 A°$$

B METALLIC RADIUS

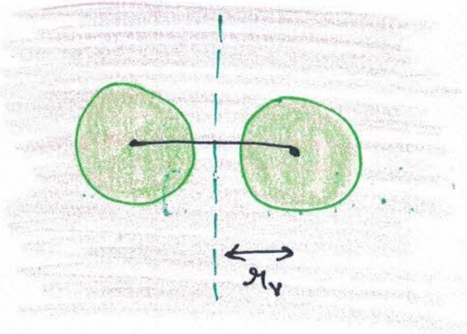


Half of the internuclear distance b/w two adjacant atoms in metallic lattice is known as metallic radius.

$$r_{metallic} > r_{covalent}$$

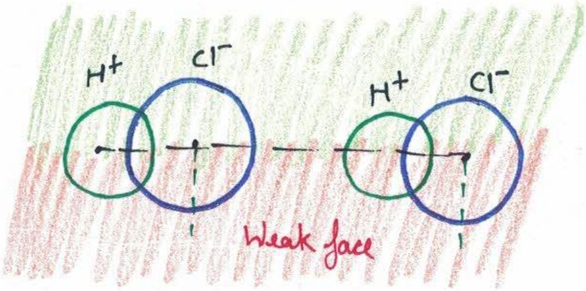
$$r_{metallic} = \frac{d_{AB}}{2}$$

C VANDER WALL'S RADIUS



Half of the internuclear distance between two adjacant atom in two adjacent molecule is known as vander wall's Radius. (in solid state).

Vander wall Radius > Metallic Radius > Covalent Radius



d IONIC RADIUS

Distance b/w nucleus and electron cloud limit is known as ionic radius.

(i) Cationic Radius

Radius of cation is smaller than radius of neutral atom.

$Na - 1e^- \longrightarrow Na^+$

$\downarrow$ K, L, M $\qquad \searrow$ K, L

Reason →
- 1 shell is removed.
- Zeff of cation > Zeff of neutral atom.

(ii) Anionic Radius

Radius of anion is larger than radius of neutral atom.

$Cl + 1e^- \longrightarrow Cl^-$

(2, 8, 7) $\qquad$ (2, 8, 8)

Reason →
- $1e^-$ repulsion will increases, size ↑
- Zeff of anion < Zeff of neutral atom

Factor Affecting Atomic Radius

(i) No. of Shells

Atomic radius $\propto$ no. of shells

down the group, as no. of shells increases, atomic radius also increases.

Trend of 1st group. (size).

H < Li < Na < K < Rb < Cs

(ii) Screening Effect

Atomic radius $\propto$ Screening effect

down the group screening effect increases, atomic radii increases.

H < Li < Na < K < Rb < Cs

(iii) Zeffective

Atomic radius $\propto \frac{1}{Zeff}$

Left to right across the period, Zeff increases atomic radii decreases.

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

(iv) Lanthanide Contraction

Atomic radii $\propto \frac{1}{\text{Lanthonide contraction}}$

As Zeff increases, size decreases

Isoelectronic species

The species containing same number of electrons are known as isoelectronic species.

$10e^-$ series ⇒ N^{-3}, O^{-2}, F^-, Ne, Na^+, Mg^{+2}, Al^{+3}

$\frac{Z}{e} \rightarrow 0.7, 0.8, 0.9, 1, 1.1, 1.2, 1.3$

Atomic radius $\propto \frac{Z}{e}$

$N^{-3} > O^{-2} > F^- > Ne > Na^+ > Mg^{+2} > Al^{+3}$

Less charage, more radius

$18e^-$ series ⇒ $Si^{-4} > P^{-3} > S^{-2} > Cl^- > Ar > K^+ > Ca^{+2} > Sc^{+3} > Ti^{+4}$

$2e^-$ series ⇒ H^-, He, Li^{+1}

$H^- > He > Li^{+1}$

PERIODIC TREND

In Periods

S and P block elements

Li > Be > B > C > N > O > F < Ne (only Zeff)

Neon has vander wall's radius

Alkari metals > Inert gas > IIIA > IVA > VA > VIA > VII A

- Alkari metals have largest size in their period.
- Inert gases have 2nd largest size in their period.
- Halogens are smallest in size in their period.
- Left to right across, period size decreases.

D block elements

$_{21}Sc > _{22}Ti > _{23}V > _{24}Cr > _{25}Mn \simeq _{26}Fe \simeq _{27}Co \simeq _{28}Ni < _{29}Cu \ll _{30}Zn$

Sc to Mn, size decreases; Mn to Ni size remains almost same and then size again increases due to inter- electronic repulsion b/w period e^-.

Mn, Fe, Co, Ni have almost same size as the screening effect and Zeff counter balance each other in this atoms.

f block elements

$_{58}Cr$ to $_{71}Lu$ size gradually decreases due to Lanthanide contraction.
$_{90}Th$ to $_{103}Lr$ size decreases due to Actinide contraction.

In Groups

S and P blocks

Atomic radii $\propto$ no. of shells
or screening constant.
down the group atomic radii increases.

IA Group	H < Li < Na < K < Rb < Cs
II A Group	Be < Mg < Ca < Sr < Ba
III A Group	B < Al < Ga < In < Tl

In Al, Z_{eff} increases as 10d elements occur before it, so its size is greater than that of Ga.

Al → 143 pm   Ga → 135 pm

IV A Group	C < Si < Ge < Sn < Pb
V A Group	N < P < As < Sb < Bi
VI A Group	O < S < Se < Te < Po
VII A Group	F < Cl < Br < I < At
0 Group	He < Ne < Ar < Kr < Xe < Rn

d block

3d < 4d $\simeq$ 5d {L.C}

3d	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn
$\overset{\frown}{4d}$	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	cd
$\overset{\simeq}{5d}$	Hf	Ta	W	Os	Re	Ir	Pt	Au	Ag

f block

4f > 5f

Actinide contraction is more powerful than lanthanide contraction.

APPLICATION OF ATOMIC RADIUS

1. LATTICE ENERGY

Energy required to break bonds in an ionic compound.

Energy released when one mole ionic compound is formed from its constituent ions is known as Lattice energy.

$Na^+ + Cl^- \longrightarrow NaCl + U$

$U \propto Z^+Z^-$ (charge on cation and anion)

$U \propto \frac{1}{r_c + r_a}$

$r_c + r_a \longrightarrow$ interionic distance

Eg.

1. $LiF > LiCl > LiBr > LiI$ (v)
2. $LiCl > NaCl > KCl > RbCl$ (u)
3. NaF MgO Na^+F^- Mg^{+2} O^{-2}

size → $Na^+ > Mg^{+2}$, $O^{-2} > F^-$

charge → $Na^{+1} < Mg^{+2}$, $O^{-2} > F^-$, $HaF > MgO$

charge is dominating

4. $NaCl$ $MgCl_2$ $AlCl_2$

size → $Na^+ > Mg^{+2} > Al^{+3}$

charge → $NaCl < MgCl_2 < AlCl_3$

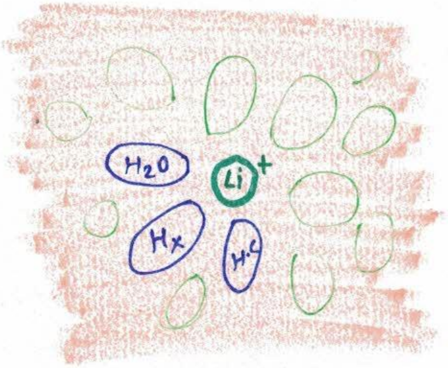
3. IONIC MOBILITY

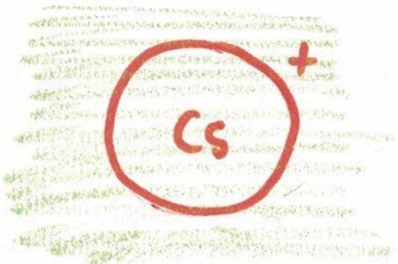
(movement of ions)

$$\text{Ionic mobility} \propto \frac{1}{\text{size of hydrated ion}}$$

$$\text{size of hydrated ion} \propto \text{hydration energy}$$

Hydrated ion → when the ion interacts with water, then forms a water layer on itself small size ion will have high hydration ionergy and big hydrated ion,





(i) Li Na K Rb Cs

Hydration energy - Li > Na > K > Rb > Cs

size of hydrated ion - $Li^+_{(aq)} > Na_{(aq)} > K^+_{(aq)} > Rb^+_{(aq)} > Cs^+_{(aq)}$

ionic mobility - $Li^+ < Na^+ < K^+ < Rb^+ < Cs^+$

(ii) Na^{+1} Mg^{+2} Al^{+3}

hydration energy - $Na^{+1} < Mg^{+2} < Al^{+3}$

size of hydrated ion - $Na^+_{(aq)} < Mg^{+2}_{(aq)} < Al^{+3}_{(aq)}$

ionic mobility - $Na^+ > Mg^{+2} > Al^{+3}$

(iii) Li^+ Be^{+2}

size of hydrated ion - $Li^+ < Be^{+2}$

ionic mobility - $Li^{+1} > Be^{+2}$

IONISATION POTENTIAL

Energy required to remove loosely held electron of outermost shell of isolated gaseous atom is known as ionisation potential or I.P.

unit = eV/atom or KJ/atom or Kcal/atom

Sublimation energy, energy required to convert 1 mole solid substance to one mole gaseous substance is known as sublimation energy.
It is an endothermic process.

$$M_{(s)} + S.E \longrightarrow M_{(g)} \quad \text{(endothermic)}$$

$$M_{(g)} + IE_1 \longrightarrow M^+_{(g)} + e^-$$

neutral atom

$$M^+_{(g)} + IE_2 \longrightarrow M^{+2}_{(g)} + e^-$$

monovalent cation

$$M^{2+}_{(g)} + IE_3 \longrightarrow M^{+3} + e^-$$

divalent cation

Increasing order of IP = $IE_1 < IE_2 < IE_3$

Reason -
- Zeff increases (nuclear charge increases)
- atomic radii decreases

Factors Affecting Ionisation Potential

1 Atomic Size

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

down the group as atomic size increases, I.P decreases.

H > Li > Na > K > Rb > Cs (I.P)

2 Screening Effect

$$I.E \propto \frac{1}{\text{screening constant } (\sigma)}$$

Screening constant (σ) down the group increases, therfore. I.P down the group decreases.

E.g. H > Li > Na > K > Rb > Cs

3 Zeff (effective nuclear charge)

$$I.P \propto Z_{eff}$$

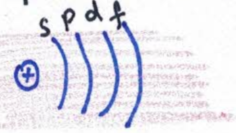
Left to Right across the period, Zeff increases.

IP increases

E.g. IP of Li < IP of Be

4. Penetration Effect

Repulsion between subshell's is penetration effect.



Penetrating power – ns > np > nd > nf

I.P ∝ Penetrating Power

E.g. IP of Be > IP of B {although Zeff Be < Zeff B}

Be B

$1s^2 2s^2$ $1s^2 2s^2 2p^1$

Penetration power of 2p < 2s

IP of IIA group > IP of IIIA group

$ns^2 > ns^2 np^1$ (I.P. order)

5. Lanthanide Contraction

I.P ∝ Lanthanides contraction

Size – 3d < 4d ≃ 5d

Correct order of I.P – 3d > 4d > 5d

4d < 3d < 5d

6. Stable Electronic Configuration

If removal of e⁻ takes place from stable e⁻ configuration, stable configuration becomes unstable, therefore more amount of energy is required, but if removal of e⁻ takes place from unstable e⁻ configuration unstable becomes stable, therefore less amount of energy required.

E.g. IP of N > IP of O

$1s^2 2s^2 2p^5$ $1s^2 2s^2 2p^4$

$\underset{1s^2 2s^2 2p^3}{N} + IE_1 \longrightarrow \underset{1s^2 2s^2 2p^2}{N^{+1}} + e^-$ ~~$1s^2$~~

stable ⟶ unstable

$\underset{1s^2 2s^2 2p^4}{O} + IE_1 \longrightarrow \underset{1s^2 2s^2 2p^3}{O^{+1}} + e^-$

unstable ⟶ stable

But

IP_2 of N < IP_2 of O

$\underset{1s^2 2s^2 2p^2}{N^+} + IE_2 \longrightarrow \underset{1s^2 2s^2 2p^1}{N^{+2}} + e^-$

unstable stable

$\underset{1s^2 2s^2 2p^3}{O^+} + IE_2 \longrightarrow \underset{1s^2 2s^2 2p^2}{O^{2+}} + e^-$

stable unstable

IP_1 of P > IP_1 of O

or IP of VA group > IP_1 of VIA group

but IP_2 of VA group > IP_2 of VIA group

PERIODIC TRENDS

Periods

S and P - block

left to right across the period, IP increases.

Reason - IP $\propto$ Zeff

IP $\propto$ Penetration power

IP $\propto$ stable e⁻ configuration

E.g. II Period

$Li < Be > B < C < N > O < F < Ne$

$\underbrace{2s^1 \quad 2s^2 \quad 2s^2 2p^1}_{\text{Penetrating power of } 2s > 2p} \quad 2s^2 2p^2 \quad \underbrace{2s^2 2p^3 \quad 2p^2 2p^4}_{\text{stable } e^- \text{ configuration}} \quad 2s^2 2p^5 \quad 2s^2 2p^6$

Least IP → alkali metal

highest IP → inert gases

d - block

3d series

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

$3d^1 4s^2$ $3d^2 4s^2$ $3d^3 4s^2$ $3d^5 4s^1$ $3d^5 4s^2$ $3d^6 4s^2$ $3d^7 4s^2$ $3d^8 4s^2$ $3d^9 4s^1$ $3d^{10} 4s^2$

Exception - IP of Ti > IP of V

IP of Co > IP of Ni

IInd IP of Cu > IInd IP of Zn

IInd IP of Cr > IInd IP of Mn

f - block

$_{58}Ce$ to $_{71}Lu$ IP gradually increases

IP $\propto$ Lanthanide contraction

Groups

S and P - block

I A $H > Li > Na > K > Rb > Cs > Fr$ → Francium is radio active

II A $Be > Mg > Ca > Sr > Ba > Ra$ → Radium is radioactive

III A B > Al > Ga > In > Tl

B > Tl > Ga > Al > In

IV A group → $C > Si > Ge > \underbrace{Sn > Pb}_{L.C.}$

C > Si > Ge > Pb > Sn

V A group → N > P > As > Sb > Bi

VI A group → O > S > Se > Te > Po

VII A group → F > Cl > Br > I > At

(0) zero group → He > Ne > Ar > Kr > Xe > Rn

d- block

$\underbrace{3d > 4d}_{A.R}\underbrace{> 5d}_{L.R}$

Correct order – 5d > 3d > 4d

f block

I.P of actinide series is more than I.P of lanthanide series due to actinide contraction

5f > 4f

APPLICATION OF IONISATION POTENTIAL

① Metallic and Non Metallic Nature

Left to Right across a period, I.P increases

non metallic nature ∝ I.P

Down the group, IP decreases, metallic nature increase

metallic nature $\propto \frac{1}{I.P}$

Q. Correct increasing order of non- metallic nature of B, Si, Al, Cl

↓ $\overset{B}{Al} < Si < Cl$ →

Cl < Si < B < Al

Cl > Si > B > Al
{non- metallic}
(metallic)

Q. Correct increasing order of metallic nature of Na, Mg, Al, K

I·P $\underset{\underset{K}{\vee}}{Na} < Mg < Al$

$K > Na > Mg > Al$
(metallic nature)

$Al < Mg < Na < K$
(amphotheric nature)

2. Oxidation State

if difference b/w 2 successive IP ≥ 16 ev, element forms stable compound in lower oxidation state.

$\underset{_{10}[Ne]3s^1}{Na} + 5.1\,ev \longrightarrow \underset{_{10}[Ne]}{Na^+} + e^-$ IP_1

$\underset{_{10}[Ne]}{Na^+} + 47.6\,ev \longrightarrow \underset{[F]}{Na^{2+}} + e^-$ IP_2

$IP_2 - IP_1 = 47.6 - 5.1 = 42.5\,ev/atom$

$IP_1 \longrightarrow IP_2$
sudden jump

The value of IP_1 and IP_2 in alkali metals has a sudden jump. Their lower oxidation state (+1) is stable.

if difference b/w 2 successive I.P < 11 ev, atom forms stable compound in higher oxidation state.

$\underset{_{10}[Ne]3s^2}{Mg} + 7.1\,ev \longrightarrow \underset{_{10}[Ne]3s^1}{Mg^{+1}} + e^-$

$\underset{_{10}[Ne]3s^1}{Mg^{+1}} + 14.6\,ev \longrightarrow \underset{_{10}[Ne]}{Mg^{+2}} + e^-$

$IP_2 - IP_1 = 14.6 - 7.1 = 7.5\,ev$

7.5 is less then 11ev, Lower oxidation state (+2) is stable.

II A group elements form stable compound in (+2) oxidation state.

if difference b/w 2 successive IP is in b/w 11ev and 16 ev, element forms stable compound in Lower and higher O.S

$\underset{3s^2 3p^1}{Al} + IP_1 \rightarrow Al^+$

$\underset{3s^2}{Al^+} + IP_2 \longrightarrow \underset{3s^1}{Al^{+2}}$

$\underset{3s^1}{Al^{+2}} + IP_3 \longrightarrow \underset{_{10}[Ne]}{Al^{+3}}$

$IP_2 - IP_1 \simeq 16\,ev \longrightarrow Al^+$ is stable

$IP_3 - IP_2 < 11\,ev \longrightarrow Al^{+3}$ is stable

Q. Four successive I.P of element X is – 50, 120, 295, 12000

$IP_2 - IP_1 = 70$, $IP_4 - IP_3 = 11705$

ⓐ. X^{+3} is stable
ⓑ. no. of valence $e^- \rightarrow 3$
ⓒ. e^- configuration $\rightarrow ns^2, np^1$

[after $3e^-$, atom might have achieved inert gas configuration]

Q.

IP_1	IP_2	IP_3	IP_4	IP_5
200	500	600	800	14600

$IP_2 - IP_1 = 300$ $IP_3 - IP_2 = 100$ $IP_4 - IP_3 = 200$ $IP_5 - IP_3 = 13800$

ⓐ X^{+4} is stable
ⓑ no. of valence $e^- = 4$
ⓒ. e^- configuration $= ns^2 np^2$

→ 190 205 800 $\underline{\underline{+2}}$

3. Reducing Nature

Reducing Power means cation forming tendency

$$\text{Reducing Power} \propto \frac{1}{I.P}$$

gas phase

$Li < Na < K < Rb < Cs$

aqueouse phase

$Na < K < Rb < Cs < Li$

Reducing Nature of elements depends on :–

1. Sublimation energy (heat of atomisation)
2. Ionisation potential
3. Hydration energy

Endothermic

exothermic

Oxidation Potential = SE + IE + hydration energy
Reduction Potential = O.P

if R.P. is –ve, then element is reducing agent.
if R.P. is +ve, then element is oxidising agent.

Li(aq) is strongest reducing agent because its heat of hydration is much higher than the sum of S.E and I.P ⇒ S.E + I.P < hyd.

F(aq) is strongest oxidising agent because its bond dissociation energy is very low (lone pair – pair repulsion occurs). Also, hydration energy of F is very high.

R.P = S.E + I.P + hyd. energy

R.P = 0 + endo. + exothermic

R.P = +ve

and R.P is +ve, F is strongest oxiding agent

4. Nature of Hydroxides

NaOH

$IP_H > IP_{Na}$

Na^+ $\underbrace{OH^-}_{\text{basic hydroxide}}$

ClOH

$IP_{Cl} > IP_H$

ClO^- $\underbrace{H^+}_{\text{acidic hydroxide}}$

($HClO_4$)
hypochlorous acid

If I.P element > I.P. of H ⟶ element forms acidic hydroxide

and if I.P element < I.P of H ⟶ element forms basic hydroxide

✓ down the group, basic nature of hydroxide increases.

$LiOH < KOH < RbOH < CsOH$ (basic nature)

Left to right, acidic nature of hydroxide increases.

$NaOH < Mg(OH)_2 < Al(OH)_3 < Si(OH)_4 < P(OH)_3 < O_2S(OH_2) < Cl(OH)$

(acidic nature)

ELECTRON .... AFFINITY

Energy released when addition of one e⁻ takes place on valence shell of a neutral isolated gaseous atom is known as electron affinity.

Ist E.A. may be +ve or −ve depending on electronic configuration of atom.

+ve Endothermic   −ve exothermic

IInd E.A is always +ve or endothermic.

Addition of 2nd e⁻ in valence shell of anion previous e⁻, repeat the incoming e⁻. To overcome interelectronic repulsion, some amount of energy is required and process becomes endothermic.

Factor Affecting Electron Affinity

1. Atomic Radius

$$E \cdot A \propto \frac{1}{\text{radius}}$$

down the group A.R. increases, E.A. decreases

I.A. ⇒ Li > Na > K > Rb > Cs

as size of atom increases, nuclear force of attraction towards incoming e⁻ decreases. Therefore, less amount of heat releases

2. Screening Effect

$$E \cdot A \propto \frac{1}{\text{radius}}$$

down the group, as A.R increase, E.A. decrease

I.A. ⇒ Li > Na > K > Rb > Cs

3. Zeff

$$E \cdot A \propto Z_{eff}$$

Left to right across the period, Zeff increases, E.A. increases.
E.A. is the property of non-metals

4. Lanthanide Contraction

$$E \cdot A \propto L \cdot C$$

5. Stable electronic Configuration

$F + 1e^- \longrightarrow F^-$ exothermic
(2,7) (2,8) stable

$Ne + 1e^- \longrightarrow Ne^-$ endothermic
(2,8) (2,8,1) unstable

$C + 1e^- \longrightarrow e^-$ exothermic
(1,4) $2s^2 2p^2$ $2s^2 2p^3$ stable

$N + 1e^- \longrightarrow N^-$ endothermic
$2s^2 2p^3$ $2s^2 2p^4$ unstable

when addition of e⁻ takes place on stable e⁻ configuration, then the element becomes unstable, less amount of heat is released or energy is absorbed to accommodate the e⁻ the process becomes endothermic.

After addition of e^- if atom gets stable configuration, more amount of heat is released and process becomes exothermic.

E.g. $x = EA_1 = -325$ which has highest
$y = EA_1 = -155$ E.A. ?
$z = EA_1 = +400$ Ans $\rightarrow x$

E.g.

A	B	C
-56	+110	+78

A > C > B
(E.A)

$-ve$ EA means easy addition of e^-
(exothermic process)

Periadic Trends in Electron Affinity

Periods

Left to right across the period, electron affinity increases.

Li	Be	B	C	N	O	F	Ne
$2s^2$	$2s^2$	$2s^2 2p^1$	$2s^2 2p^2$	$2s^2 2p^3$	$2s^2 2p^4$	$2s^2 2p^5$	$2s^2 2p^6$
$\downarrow +1e^-$	$\downarrow +1e^-$	$\downarrow +1e^-$	$\downarrow +1e^-$	$\downarrow +1e^-$	$\downarrow +1e^-$	$\downarrow +1e^-$	$\downarrow +1e^-$
Li^-	Be^-	B^-	C^-	N^-	O^-	F^-	Ne^-
$2s^2$	$2s^2 2p^1$	$2s^2 2p^2$	$2s^2 2p^3$	$2s^2 2p^4$	$2s^2 2p^5$	$2s^2 2p^6$	
(exo)	(endo)	(exo)	(exo)	(endo)	(exo)	(exo)	(endo)

Ne < Be < N < Li < B < C < O < F
(new shell) (new subshell)

- Highest electron affinity = Halogens
- Lowest electron affinity = Inerst gas

Order of E.A.

inert gas < alkaline < earth metal < VA < alkali metal < IIIA < IVA < VIA < Halogens

Inert gases have lowest EA
and highest IP

Groups

down the group, E.A. decreases

$$E.A. \quad \frac{1}{\text{Atomic radii}}$$

IA group H Li Na K Rb Cs

II A group Be Mg Ca Sr Ba

III A to VIII A

EA of 3rd period > EA of 2nd period

V A Nitrogen family

$N \underset{\text{endo}}{} < \underset{\text{exo}}{P} > As > Sb > Bi \Rightarrow P > N > As > Sb > Bi$

size of 3P > size of 2P

VI A oxygen family

O < S > Se > Te > Po

order:- S > Se > Te > Po > O

VII A Halogens

F < Cl > Br > I

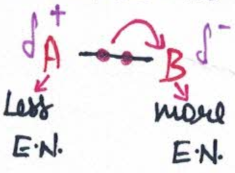
order:- Cl > F > Br > I

chlorine has highest electron affinity in the periodic table.

ELECTRONEGATIVITY

Electron attracting power of shared pair of e⁻ of atom A or B towards itself covalent bonded molecule is known as electronegativity.

E.g.



Electronegativity is a relative value. It is dimensionless.

Factors Affecting Electronegativity

Atomic radii

$$E.N \propto \frac{1}{\text{radii}}$$

Down the group, AR increases, EN decreases.

$H > Li > Na > K > Rb > Cs$

Left to right EN increases

$Li < Be < B < C < N < O < F$

E.N. of inert gas = 0

2. Zeff

E.N ∝ Zeff

Left to right, Zeff ↑ E.N ↑

$Na < Mg < Al < Si < P < S < Cl$

3. Oxidation state

E.N ∝ +ve O.S

$Fe^{+3} > Fe^{+2} > Fe^{0}$

$Sn^{+2} > Sn^{0}$

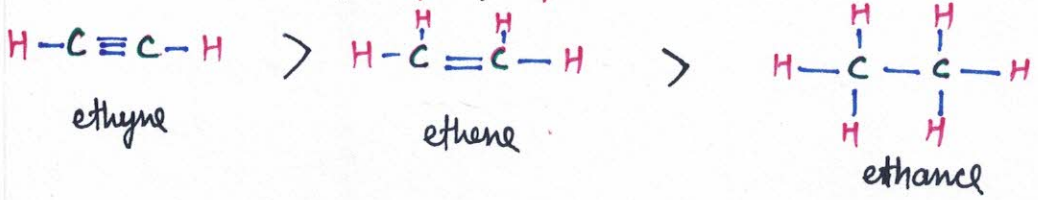
$Na^{+} > Na^{0}$

4. Percentage and character in hybrid orbitals

E.N ∝ % s character in hybrid orbitals

SP	sp²	sp³
$\frac{1}{2} \times 100$	$\frac{1}{3} \times 100$	$\frac{1}{4} \times 100$
50%	33.3%	25%

$SP > sp^2 > sp^3$



ELECTRONEGATIVITY SCALE

1. Pauling's Scale

Pauling preferred bond energy of a 100% covalently bonded molecule to calculate the E.N.

He assumed E.N of flourine as 4.

$H_2 + F_2 \longrightarrow 2HF$

$$H-H \quad + \quad F-F \xrightarrow{\quad} 2H-F$$

$$\underset{x}{H-H} \qquad \underset{y}{F-F} \xrightarrow{\quad} \underset{z}{H-F}$$

Calculate bond energy $= \frac{1}{2}x + y$ or

Geometric mean $= \sqrt{xy}$

ΔAB = change in bond energy

ΔAB = BE observed − BE calculated

Resonance energy $= z - \sqrt{xy}$

$$0.208\sqrt{\Delta AB} = (x_A - x_B)$$

x_A → electronegativity of more electronegative element

x_B → electronegativity of less electronegative element

E.g. $^{x}E_{H-H} = 104.2$ Kcal mol^{-1}

$^{Y}E_{H-H} = 36.6$ Kcal mol^{-}

$Z = E_{H-F} = 134.6$ Kcal mol^{-}

$X_H = 2.1$

$\{\Delta AB = z - \sqrt{xy}$

$$\Delta AB = 134.6 - \sqrt{104.2 \times 36.6}$$

$$\Delta AB = 134.6 - 62$$

$$\Delta AB = 72.6$$

$$0.208\sqrt{\Delta AB} = (X_F - X_H)$$

$$0.208\sqrt{72.6} = (X_F - 2.1)$$

$$0.208 \times 8.5 = X_F - 2.1$$

$$1.7 = X_F - 2.1$$

$$1.7 + 2.1 = X_F$$

$$X_F = 3.8 \simeq 4$$

2. Mulliken's Scale

It is based on I.P and E.A.

$$X_M = \frac{I.P + E.A}{2}$$

$$X_P = \frac{X_M}{2.8}$$

$$X_M = \frac{IP + EA}{5.6} \text{ ev atom}^{-}$$

1eV = 23.06 Kcal

$$X_M = \frac{IP + EA}{2 \times 62.5} \text{ Kcal mol}^{-1}$$

$$X_M = \frac{IP + EA}{125} \text{ Kcal mol}^{-1}$$

Q. I.P and E.A of an element A are 13.7 and 3.8 ev/atom respectively what is electronegativity of element A ?

A.

$$X_M = \frac{IP + EA}{5.6} \text{ ev/atom}$$

$$= \frac{13.7 + 3.8}{5.6} = \frac{17.5}{5.6} = 3.125$$

Q. IP = 400 Kcal EA = 80 Kcal

$$EN = \frac{400 + 80}{125} = \frac{480}{125} = \frac{96}{25} = 3.9$$

$$\left\{ EN = \frac{IP + EA}{125} \text{ Kcal mol}^{-1} \right\}$$

It is based on Zeff

3. Alfred Rachow's scale

$$X_{AP} = 0.359 \frac{Z_{eff}}{r^2 (\text{in A}^\circ)}$$

r → radius of atom in Angstrons

$$X_{AR} \rightarrow 0.744 + \frac{0.359\ Z_{eff}}{r^2}$$

Q. $r_{si} = 1.175\ A^\circ$ E.N. $_{Si}$ = ?

$$Z_{eff} = 2.2 * 3 \times 65 = 6.6 * 0.65 = 4.290 = 4.19$$

$$X_{Si} = 0.744 + \frac{0.359 \times 4.19}{1.175 \times 1.175}$$

$$X_{Si} = 0.744 + 0.2$$

$$= 0.944 \simeq 1$$

PERIODIC TRENDS

Left to right across the period, electronegativity increases.

$EN \propto Z_{eff}$

Down the group, EN decreases

$EN \propto \frac{1}{size}$

I	H 2.1							He 0
II	Li 1.0	Be 1.5	B 2	C 2.5	N 3	O 3.5	F 4	Ne 0
III	Na 0.9	Mg 1.2	Al 1.5	Si 1.8	P 2.2	S 2.5	Cl 3	Ar 0
	K 0.8						Br 2.8	
	Rb 0.8						I 2.5	
	Cs 0.7							

E.N of inert gas is 0

L to R II period, EN increases by 0 unit

L to R in III period, EN increases by 0.3 unit

Exception :- Zn < Cl < Hg [EN down the group ↑↑]

APPLICATION OF ELECTRONEGATIVITY

1. Metallic & Non Metallic Nature

non metallic nature $\propto$ EN

metallic nature $\propto \frac{1}{EN}$

down the group, metallic nature increases & left to right across a period, non-metallic nature increases.

Q. Write the correct order of metallic nature :- B, C, Si, Na, K

A. $\downarrow \begin{matrix} Na \\ K \end{matrix} \longrightarrow \begin{matrix} B & C \\ & Si \end{matrix} \downarrow$

Order:- C < B < Si < Na < K

2. Polar and Non-Polar Bonds

Non Polar Bond → H–H   Cl–Cl   N–N

Polar Bond → O–H   N–H   Cl–H

$$\text{Bond Polarity} \propto \Delta EN$$

$$\Delta EN = X_A - X_B$$

Q. Which bond is most polar

(a) C–H   (b) N–H   (c) O–H   (d) N–N

$\Delta = 0.3$ $\Delta = 0.9$ 1.4   3-3=0

A. most polar = O–H

least polar = N–N

3. Nature of Hydroxides

$\underbrace{Na}_{1.0}\underbrace{O}_{3.5}\underbrace{H}_{2.1}$

$En\,|Na - O| = 2.5 \longrightarrow Na^+$ Polar bond

OH^- basic nature

X_o - element > X_o - Hydrogen

O → oxygen then, the hydroxide is basic

Cl OH

3.0 3.5 1

$En\,|O - Cl| = 0.5$

$En\,|O - H| = 1.4$

$En_{O-Cl} < En_{O-H} \longrightarrow$ acidic nature

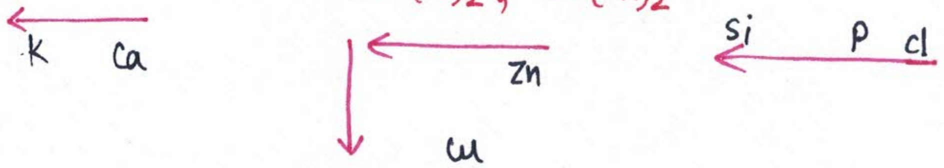
then the hydroxides is acidic.

L to R acidic nature ↑

Top to bottom basic nature ↑

Q. Arrange in basic order:- KOH, $Ca(OH)_2$, $Si(OH)_4$, $P(OH)_3$, $ClO_3(OH)$, $Cu(OH)_2$, $Zn(OH)_2$

A.



Cl < P < Si < Zn < Cu < Ca < K

HClOH < H_3PO_4

$ClO_3(OH) < PO(OH)_3 < Si(OH)_4 < Zn(OH)_2 < Cu(OH)_2 < Ca(OH)_2 < KOH$

4. Nature of Oxides

Left to right across the period, acidic nature of oxides increases, down the group, basic nature of oxides increases, Generally metallic oxides are basic oxides, metalloid oxides are amphotheric and non-metallic oxides are acidic oxides.

amphotheric elements ⇒ Zn, Al, Be, Ga, Pb, Sn, As, Sb, Bi, Ge

Q. No. of amphotheric oxide acidic and basic

i) N_2O_5 ii) SO_2 iii) Al_2O_3 iv) GeO_3 v) Sn_2O_3

vi) ZnO vii) K_2O viii) MnO_2

A.
Amphotheric - 5 Al_2O_3 GeO_2 SnO_3 ZnO MnO_2
acidic - 2 N_2O_5 SO_2
Basic - 1 K_2O

when intermediate o.s of d-block elements amphoteric oxide will be formed.

E.g. Amphoteric oxides → Cr_2O_3, MnO_2 UO_2

as % of oxygen ↑ acidic nature increases

as +ve oxidation state increases acidic nature increases.

$$\underbrace{\overset{MnO}{+2} \quad \overset{Mn_2O_3}{+3}}_{basic} \quad \underbrace{\overset{MnO_2}{+4}}_{ampotheric} \quad \underbrace{\overset{Mn_2O_5}{+5} \quad \overset{MnO_3}{+6} \quad \overset{Mn_2O}{+7}}_{Acidic}$$

5. Bond Length

Internuclear distance b/w the covalent bonded atoms is known as bond length.

$$d_{AB} = r_A + r_B - 0.09\,|x_A - x_B|$$

$$\text{bond length} \propto \frac{1}{\Delta En}$$

1. C–Cl < C–Br < C–I
2. C–H > N–H > O–H

6. % age Ionic character in Compound

Harry Smith's formula

$$\%\ \text{ionic character} = 16\,[x_A - x_B] + 3.5\,[x_A - x_B]^2$$

E.g. CsF

$$\%\ \text{ionic character} = 16\,[4-0.7) + 3.5\,[4-0.7]$$

$$= 16 \times 3.3 + 3.5 \times (3.3)^2$$

$$= 51.8 + 3.5 \times 10.89$$

$$= 51.8 + 38.115$$

$$= 89.915$$

$$= 90\%$$

E.g. H–F

$$= 16[4-2.1] + 3.5\,[4-2.1]^2$$

$$= 16\,[1.9] + 3.5\,(1.9)^2$$

$$= 30.4 + 3.5 \times 36.1$$

$$= 30.4 + 12.735$$

$$= 43.135\%$$

E.g. H–Cl

$$= 16\,[3-2.1] + 3.5\,(3.21)^2$$

$$= 16 \times 0.9 + 3.5 \times (0.9)^2$$

$$= 14.4 + 3.5 \times 0.81$$

$$= 14.4 + 2.835$$

$$= 17.235\%$$

Eg. H—I

$= 16[0.4] + 3.5(0.4)^2$

$= 6.4 + 3.5 \times 0.16$

$= 6.4 + 0.56$

$= 6.96$

$= 6.96\%$

ionic nature $\propto$ ΔEN

down the group ionic nature decreases and covalent nature increases.

7. Heat the Formation

change in enthalpy when 1 mole compound is formed from their constituent elements in standard state is known as Heat of formation.

$$C_{(g)} + O_{2(g)} \longrightarrow CO_{2(g)}$$

$$\Delta_f H = 23\Sigma[X_A - X_B]^2 - 56n_N - 26n_O$$

(K cal mol⁻¹)

where,

Σ = no. of bonds

n_N = no. of N. atoms

n_O = no. of O atoms

E.g.

① CO_2

O=C=O

$\Sigma = 4$ $X_C = 2.5$ $X_O = 3.5$

$\Delta H_f CO_2 = 23 \times 4[3.5 - 2.5]^2 - 56 \times 0 - 26 \times 2$

$= 92 \times 1 - 52$

$= 40$ K cal mol⁻¹

② $BeCl_2$

Cl—Be—Cl

$\Sigma = 2$ $X_{Cl} = 3$ $X_{Be} = 1.5$

$\Delta_f H_{BeCl_2} = 23 \times 2[3 - 1.5]^2 - 56 \times 0 - 26 \times 0$

$= 46(1.5)^2$

$= 46 \times 2.25$

$= 9.3$ K cal mol⁻¹

8. Acidic Nature of Oxyoxides

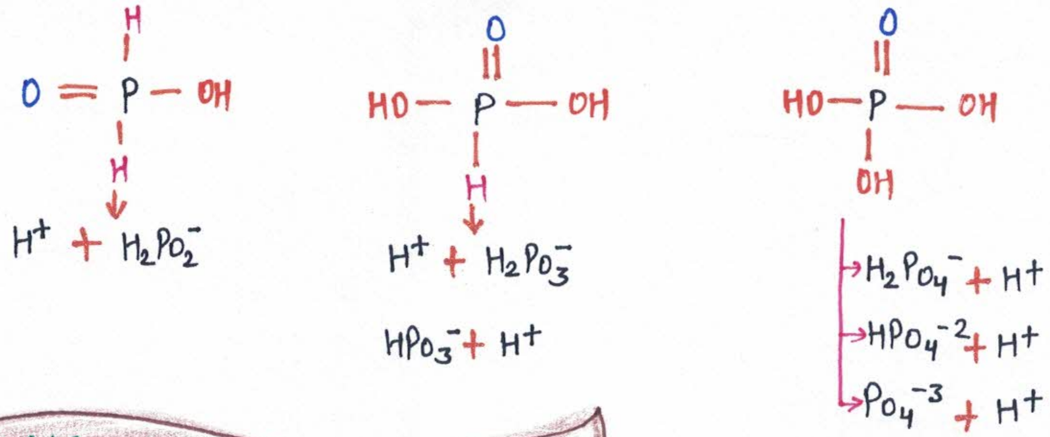
Acidic nature ∝ EN of central atom

1.	$HClO_4$ (+7)	H_2SO_4 (+6)	HNO_3 (+5)	[EN +ve O.S]
2.	$HClO$ (+1)	$HClO_2$ (+2)	$HClO_3$ (+3)	$HClO_4$ (+4)
3.	HNO_2 (+3)	HNO_3 (+5)		
4	H_2SO_3 (+4)	H_2SO_4 (+6)		

As % of oxygen increases, acidic nature of oxyacids increases.

Exception

$H_3PO_2 > H_3PO_3 > H_3PO_4$



9. Acidic Nature of Hydra Acid

(i) L to R across a period acidic nature of hydra acid increases

$NH_3 < H_2O < H-F$

basic neutral acidic

Acidic natural ∝ EN of central atom

(ii) down the group, acidic nature of hydra acid increases.

Exception:- of EN

size factor is dominating here

a. $H-F$ < $H-Cl$ < $H-Br$ < $H-I$

$H-I$ bond is weak bond, so it will furnish~~ed~~.
H^+ ion more readily in its aqueous solⁿ.

b. H_2O < H_2S < H_2Se < H_2Te **Acidic Nature**

10. Nomenclature of Inorganic Compounds

OF_2	oxygen di flouride
HF	hydrogen flouride
H_2O	hydrogen oxide
O_2F_2	oxygen monoflouride
Cl_2O_2	chlorine oxide

Exception:- NH_3 – ammonia (not hydrogen nitric)

Inert Pair Effect

Group	13	14	15
	$ns^2 np^1$	$ns^2 np^2$	$n^2 np^3$
	B +3+1	C +2+4	N +3+5
	Al	Si	P
	Ga	Se	As
	In	Sn	Sb
	Tn +1	Pb +2	Bi +3

In 13th, 14th and 15th group elements, down the group, heavier element forms stable compound in lower oxidation state because ns^2 e^- does not participate in chemica reaˣⁿ only np e^- participate in bonding. This is known as inert Pair effect

13th Group

stability of +3 O.S.

$BX_3 > AlX_3 > GaX_3 > InX_3 > TlX_3$

stability of +1 O.S

$BCl \ll AlCl < GaCl < InCl \ll TlCl$

TlX_3 act as an oxidising agent

$$TlX_3 \longrightarrow Tl + X_2$$

$$\underset{(+3)}{TlCl_3} \longrightarrow \underset{(+1)}{TlCl} + Cl_2$$

Q. What is O.S of Tl in TlI_3 ?

Ans $+1$, $TlC_3 \rightarrow Tl^{+1} + I_3^{-1}$ (asize)

Tl^{+3} I^- doesnot exist

It exist in form of $Tl^+ I_3^{-1}$

14th Group Carbon family

+4 O.S.

$CX_4 > SiX_4 > GeX_4 > SnX_4 > PbX_4$

PbX_4 acts as oxidising agent

$\underset{\text{unstable}}{PbX_4} \rightarrow \underset{\text{stable}}{PbX_2} + X_2$

Q. Why Sn^{+2} acts as a reducing agent and Pb^{+4} act as oxidising agent.

A: $Sn^{+2} - 2e^- \rightarrow Sn^{+4}$ (Reducing agent)

$Pb^{+4} + 2e^- \rightarrow Pb^{+2}$ (oxidising agent)

+2 O.S

$CX_2 < GeX_2 < SnX_2 \ll PbX_2$

15th Group

$(ns^2np^3) \rightarrow +3, +5$

+3 O.S

$NCl_3 < PCl_3 < AsCl_3 < SbCl_3 \ll BiCl_3$

+5 O.S

$NCl_5 > PCl_5 > AsCl_5 > SbCl_5 \gg BiCl_5$

Q. Which of the following order is/are correct for stable O.S.

a. $Tl^{+3} > Tl^{+1}$ b. $Tl^{+3} < Tl^{+1}$ c. $Pb^{+2} > Pb^{+4}$

d. $Sn^{+2} > Sn^{+4}$ e. $Sn^{+4} > Sn^{+2}$ f. $Be^{+3} > Be^{+5}$

A. b, c, e, f

Q. Which of the following ion act as O.A. ?

a. Tl^{+3} b. Pb^{+4} c. Be^{+5} d. Pb^{+2}

e. Sn^{+2} f. Sn^{+4}

A. a, b, c

DIAGONAL RELATION-SHIP

II Period :-	Li 1	Be 1.5	B 2	C	N	O	F	Ne
III Period :-	Na 0.9	Mg 1.2	Al 1.5	Si	P	S	Cl	As

Pairs :- Li-Mg Be-Al B-Si

Reason :- ① ionic potential = $\phi = \frac{charge}{size}$

② atomic size (ionic size

③ Electronegativity

E.g. Li $0.72 A^\circ$ Mg^{+2} $0.69 A^\circ$

1. Li and Mg

Lithenium and Magnesium shows similar property Reasons:-

(i) Almost same ionic size

(ii) $EN_{Li} = 1.0$ $EN_{Mg} = 1.2$

- Li and Mg carbonate, bicarbonates, bisulphits sulphats, nitrates, hydroxides are unstable.

$Li_2CO_3 \longrightarrow Li_2O + CO_2$

$MgCO_3 \longrightarrow MgO + CO_2$

- Li and Mg both elements directly reacts with N_2 to form nitrides

$6Li + N_2 \longrightarrow 2Li\ N$

$Mg + N_2 \longrightarrow Mg_3N_2$

- Li and Mg directly reacts with carbon to form carbides.

$2Li + C_2 \longrightarrow Li^+\ \bar{C} \equiv \bar{C} Li^+$

Lithenium carbide

$2Mg + 3C \longrightarrow Mg_2C_3$

$(Mg^{+2}\ C^{-2} = C = C^{-2}\ Mg^{+2})$

Magnesium carbide

- Li and Mg forms covalent compounds. covalent character develop in comp.

2. Be and Al

Reason:- (i) Ionic Potential

(ii) $EN_{Be} = 1{\cdot}5 = EN_{Al}$

Properties:-

Be and Al both are amphoteric metal

BeO and Al_2O_3 both are amphoteric metal oxides.

ⓐ $BeO + 2HCl \longrightarrow BeCl_2 + H_2O$

basic

ⓑ $BeO + 2NaOH \longrightarrow ?$

$BeO + H_2O \longrightarrow Be\cancel{(OH)}_2$

$H_2B\cancel{eO_2} + 2NaOH \longrightarrow Na_2BeO_2 + 2H_2O$

$$\boxed{BeO + 2NaOH \longrightarrow Na_2BeO_2 + H_2O}$$

acidic

ⓒ $Al_2O_3 + 6HCL \longrightarrow 2AlCl_3 + 3H_2O$

basic

$Al_2O_3 + 6NaOH \longrightarrow ?$

$Al_2O_3 + 3H_2O \longrightarrow \cancel{Al(OH)_3}$

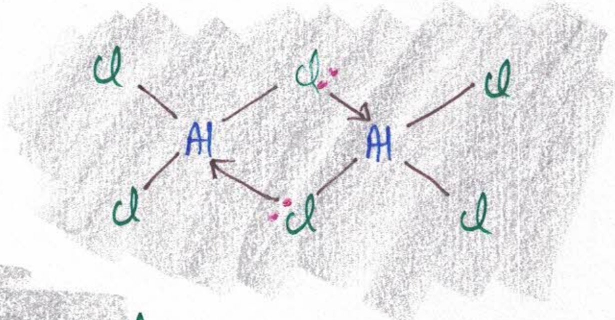
$2H_3\cancel{AlO_3} + 6NaOH \longrightarrow 2Na_3AlO_3 + 6H_2O$

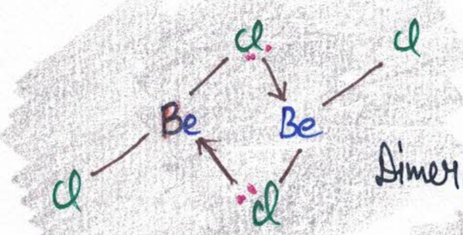
$$\boxed{Al_2O_3 + 6NaOH \longrightarrow 2Na_3AlO_3 + 3H_2O}$$

Acidic

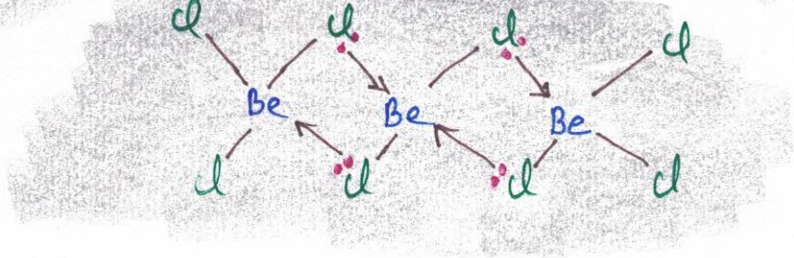
- $Be(OH)_2$ and $Al(OH)_3$ both are amphoteric hydroxides.

$BeCl_2$ and $AlCl_3$ both exist in dimer form.





Dimer(gas)



Polymer (solid)

Be and Al both are insoluble in NH_3. Any metal when dissolved in NH_3 give blue colour but Be and Al get solvated with NH_3, but does not give blue colour.

$$\left.\begin{array}{l} Be + \text{Conc. } HNO_3 \longrightarrow BeO \\ Al + \text{Conc. } HNO_3 \rightarrow Al_2O_3 \end{array}\right] \text{Passive Layer}$$

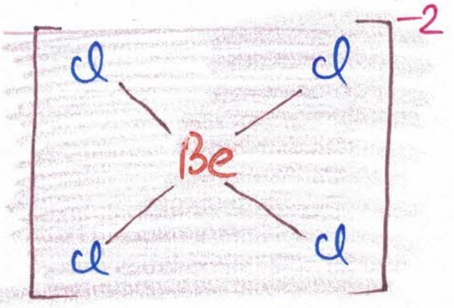
They, on reacting with HNO_3 forms a protective Layer which does not allow contact with air or moisture.

Be and Al forms complex compounds

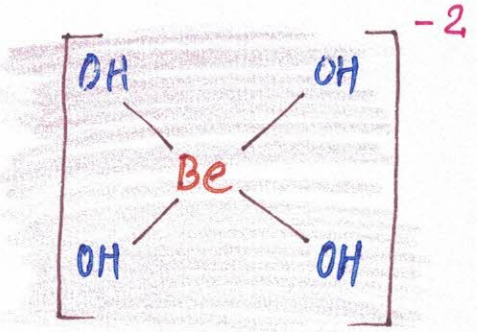
Be and Al have vacant orbitals

Be forms tetrahedral complex.

E.g.

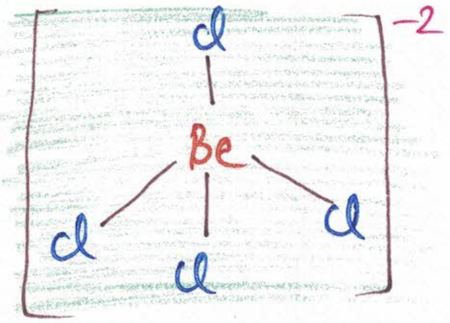


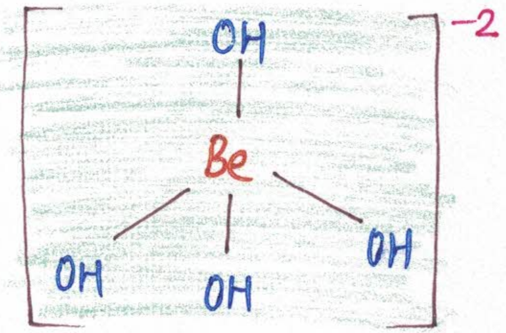
or



$(BeCl_4)^2$ or $[Be(OH)_4]^{-2}$

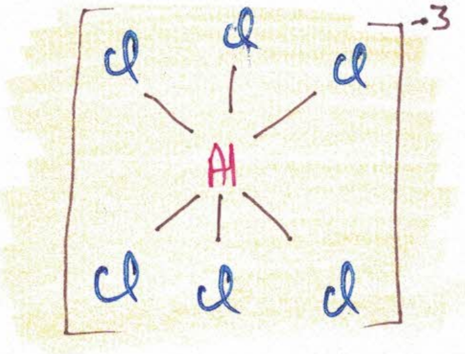
Tetrahedral





Al forms octahedral complex

E.g. $[AlCl_6]^{-3}$ or $[Al(OH)_6]^{-3}$



Reducing power of Be and Al is very Less.

R.P = SE + IE + hyd. energy

3. B and Si

Reason:- (i) ionic potential of B and Si are almost similar
(ii) $EN_B = 2$ $EN_{Si} = 1.8$

Properties:-

B and Si both are metalloids
B and Si form hard compounds.

hard Compound

- B_4C → boron carbide
  Very hard, even comparable to diamond
- BN → boron nitride
- SiO_2 → silica
- SiC → carborundum

Silicates

B and Si both dissolve in $NaOH$

$$2B + \cancel{6H_2O} \longrightarrow \cancel{H_3BO_3} + 3H_2$$
$$\cancel{2H_3BO_3} + 6NaOH \longrightarrow 2Na_3BO_3 + \cancel{6H_2O}$$
$$\boxed{2B + 6NaOH \longrightarrow 2Na_3BO_3 + 3H_2 \uparrow}$$

sodium borate

$$Si + \cancel{2H_2O} \longrightarrow SiO_2 + 2H_2$$
$$SiO_2 + H_2O \longrightarrow H_2SiO_3$$
$$\cancel{H_2SiO_3} + 2NaOH \longrightarrow Na_2SiO_3 + \cancel{2H_2O}$$
$$\boxed{Si + H_2O + 2NaOH \longrightarrow SiO_2 + Na_2SiO_3 + 2H_2}$$

sodium silicate