

Chemistry - I

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Chemistry

Inorganic

Organic

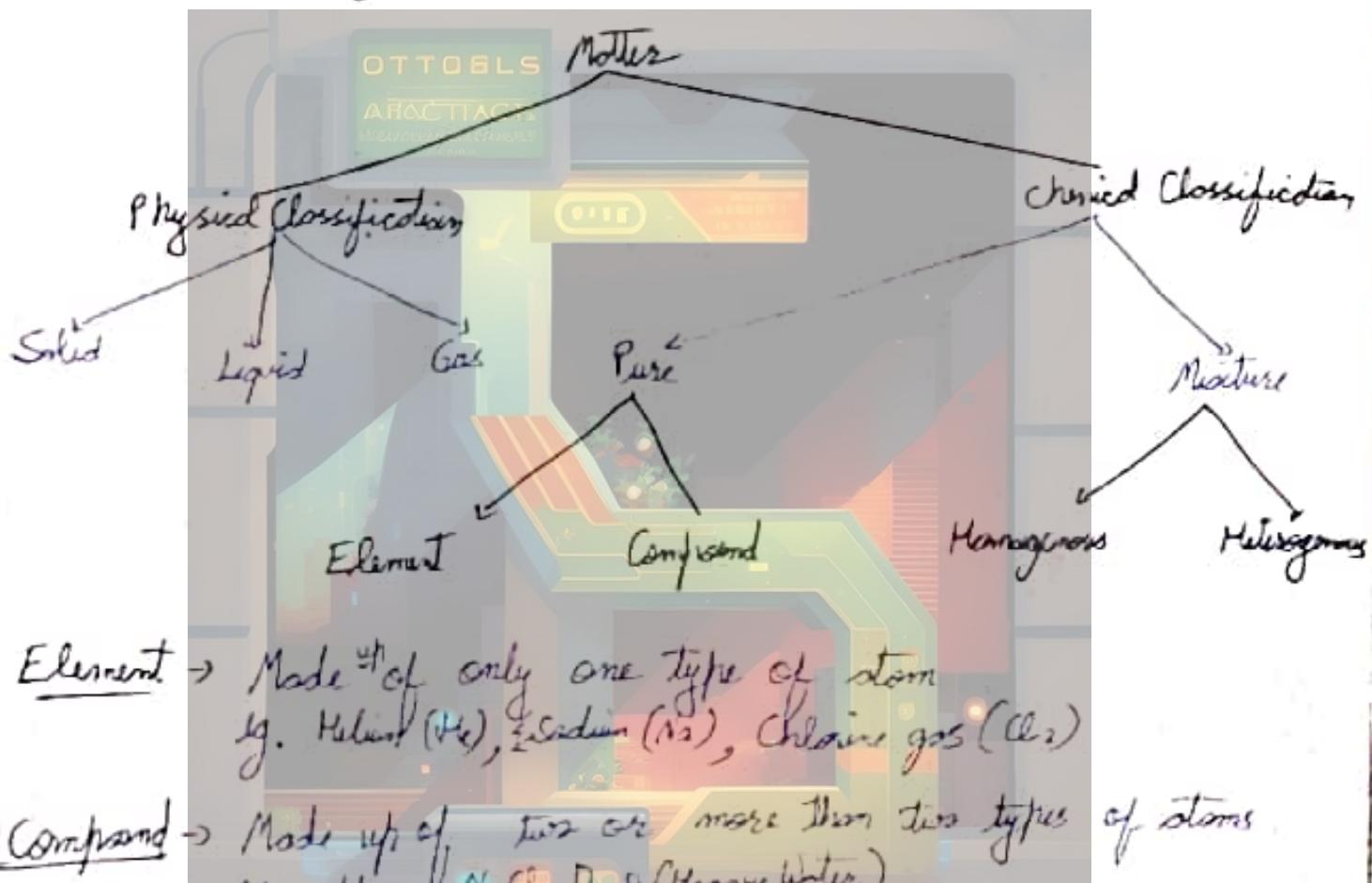
* Chemistry is a branch of science that deals with the study of matter around us.

Physical

Inorganic

Quantum Numbers & Electronic configuration

Matter → Anything which has mass & occupies some space



Element → Made up of only one type of atom
e.g. Helium (He), Sodium (Na), Chlorine gas (Cl_2)

Compound → Made up of two or more than two types of atoms
e.g. H_2O , NaCl , D_2O (Heavy Water)

→ Compounds follow fixed mass proportion

Homogeneous Mixture → It is a mixture of two or more types of substance having a uniform composition throughout

eg. Sugar + H_2O
 $NaCl + H_2O$
 Alcohol + Water
 Alloys - Brass ($Cu + Zn$)
 Bronze ($Cu + Sn$)

Heterogeneous Mixture \rightarrow It is a mixture of two or more types of substances having non-uniform composition.

eg Oil + Water
 Sand + Water
 Sand, milk, Blood

Q Classify:-

Element	Compound	Homogeneous	Heterogeneous
Gold	Glucose	Pure Air	Air with Dust
Silver	$C_{12}H_{22}O_{11}$	$NaCl + H_2O$	Ink + Water

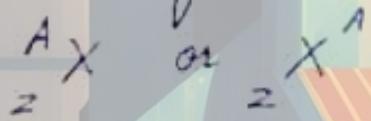
- Atom \rightarrow ① Smallest unit of matter
 ② It may or not exist in free state in nature
 eg Na, Cl
 ③ It is electrically neutral
 ④ Made up of 3 fundamental particles - electrons
 - protons
 - neutrons

Name	charge	mass	charge	location
1. Electron	$-1.602 \times 10^{-19} C$ or $-4.8 \times 10^{-10} e.s.u.$	$9.11 \times 10^{-28} g$	Negative	Outside the Nucleus
2. Proton	$+1.602 \times 10^{-19} C$ or $+4.8 \times 10^{-10} e.s.u.$	$1.6725 \times 10^{-24} g$ $\frac{1}{1837}$ mass of C	Positive	Inside the Nucleus
3. Neutron	0	$1.675 \times 10^{-24} g$	Neutral	Inside the nucle

6 e.s.u - electrostatic unit of charge

$$1.67 \times 10^{-24} \text{ gram} = 1 \text{ e.s.u}$$

Representation of a atom



A - Atomic mass number

Z - atomic number

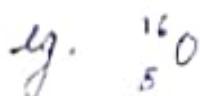
$$\text{n.z. of } {}_{+1}^1 H = 2$$

so if $e = 1 = 2$ (In a state absence of charge)

so if $e = 2$ - charge (for cation)

$2 +$ charge (for anion)

$$\text{atomic mass (A)} = n + p$$



O - symbol of oxygen atom

16 - A - mass

8 - Z - atomic no.

$$n = Z = 8$$

$$e = p = 8$$

$$n = A - Z = 16 - 8$$

$$= 8$$

$$\text{Ex - 2} \quad \text{CH}_4 \quad f_1 = 6 + 4 = 10$$

$$e = 6 + 4 = 10$$

$$\pi = (12 - 6) + 4(1-1) = 6 + 0 = 6$$

$$\text{H}_2\text{O} \quad f_1 = 2 + 8 = 10$$

$$e = 2 + 8 = 10$$

$$\pi = 0 + 8 = 8$$

$$\text{H}_2\text{O}_2 \quad f_1 = 2 + 16 = 18$$

$$e = 2 + 16 = 18$$

$$\pi = 0 + 16 = 16$$

$$\text{N}_2\text{O} \quad f_1 = 2 + 8 = 10$$

$$e = 2 + 8 = 10$$

$$\pi = 2 + 8 = 10$$

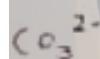
ATRACTA:



$$f_1 = 7 + 3 = 10$$

$$e = 7 + 3 = 10$$

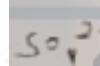
$$\pi = 7 + 0 = 7$$



$$f_1 = 6 + 24 = 30$$

$$e = 6 + 24 + 2 = 32$$

$$\pi = 6 + 24 = 30$$



$$f_1 = 16 + 32 = 48$$

$$e = 16 + 32 + 2 = 50$$

$$\pi = 16 + 32 = 48$$



$$f_1 = 16$$

$$e = 16$$

$$\pi = 16$$

+1 or 1⁺

+2 or 2⁺

+3 or 3⁺

-3 or 3⁻

-2 or 2⁻

-1 or 1⁻

→ monovalent cation

→ Divalent cation

→ Trivalent Cation

→ Tetravalent cation

→ Divalent anion

→ monovalent anion

1 - mono

2 - di

3 - tri

4 - tetra

5 - penta

6 - hexa

7 - hepta

8 - octa

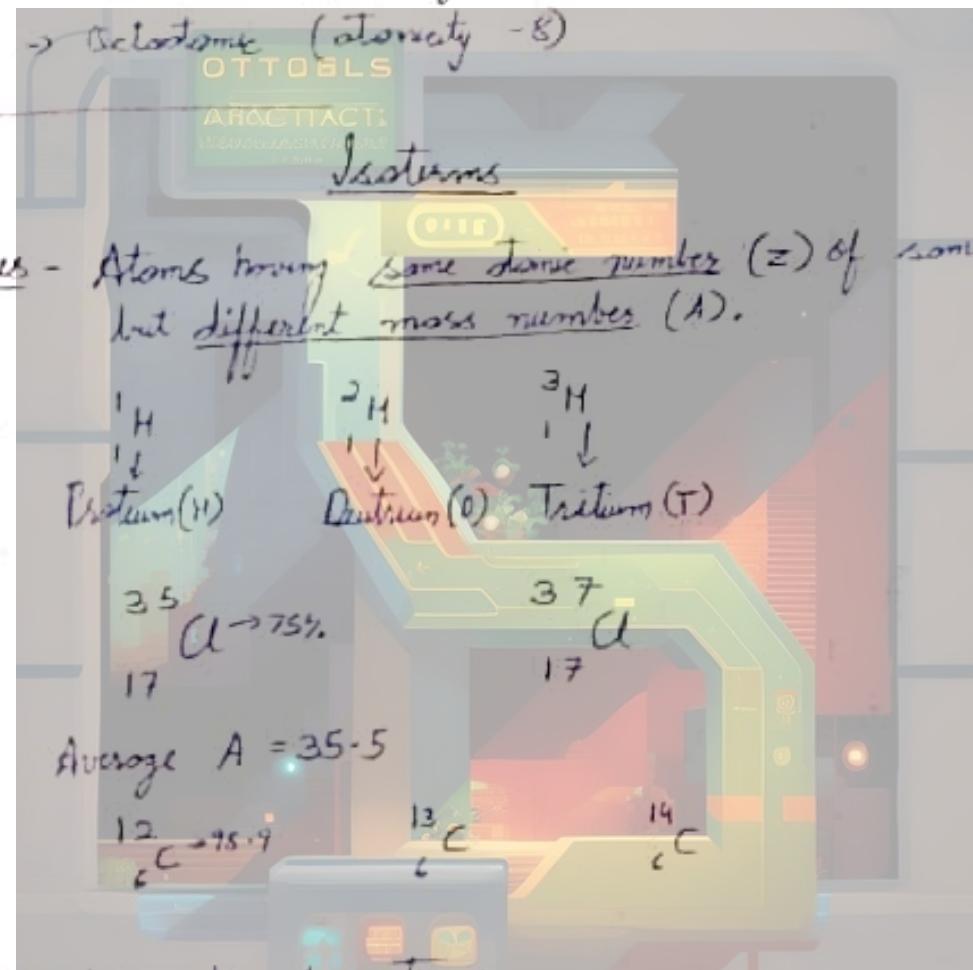
9 - nona

10 - deca

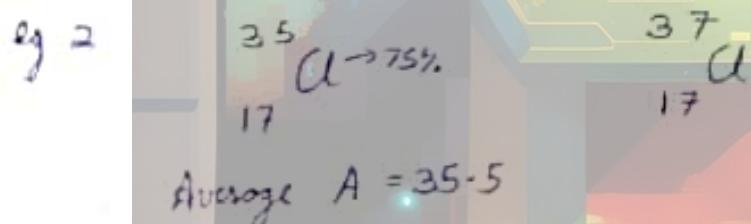
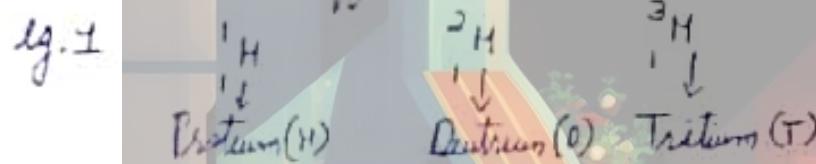
Molecule → A molecule is formed when two or more than two atoms are chemically combined (either same or different) by a covalent bond.

Atomicty → Total number of atoms present in a molecule is called its atomicty.

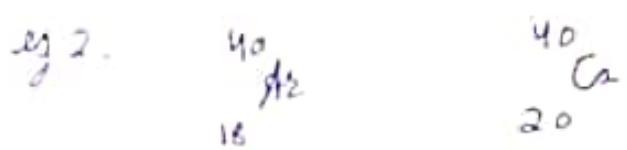
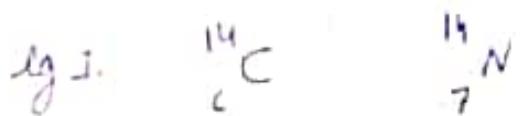
- e.g. Ne → Monatomic (atomicty -1)
 O_2 → Diatomic (atomicty -2)
 O_3 → Triatomic (atomicty -3)
 P_4 → Tetraatomic (atomicty -4)
 S_8 → Octaatomic (atomicty -8)



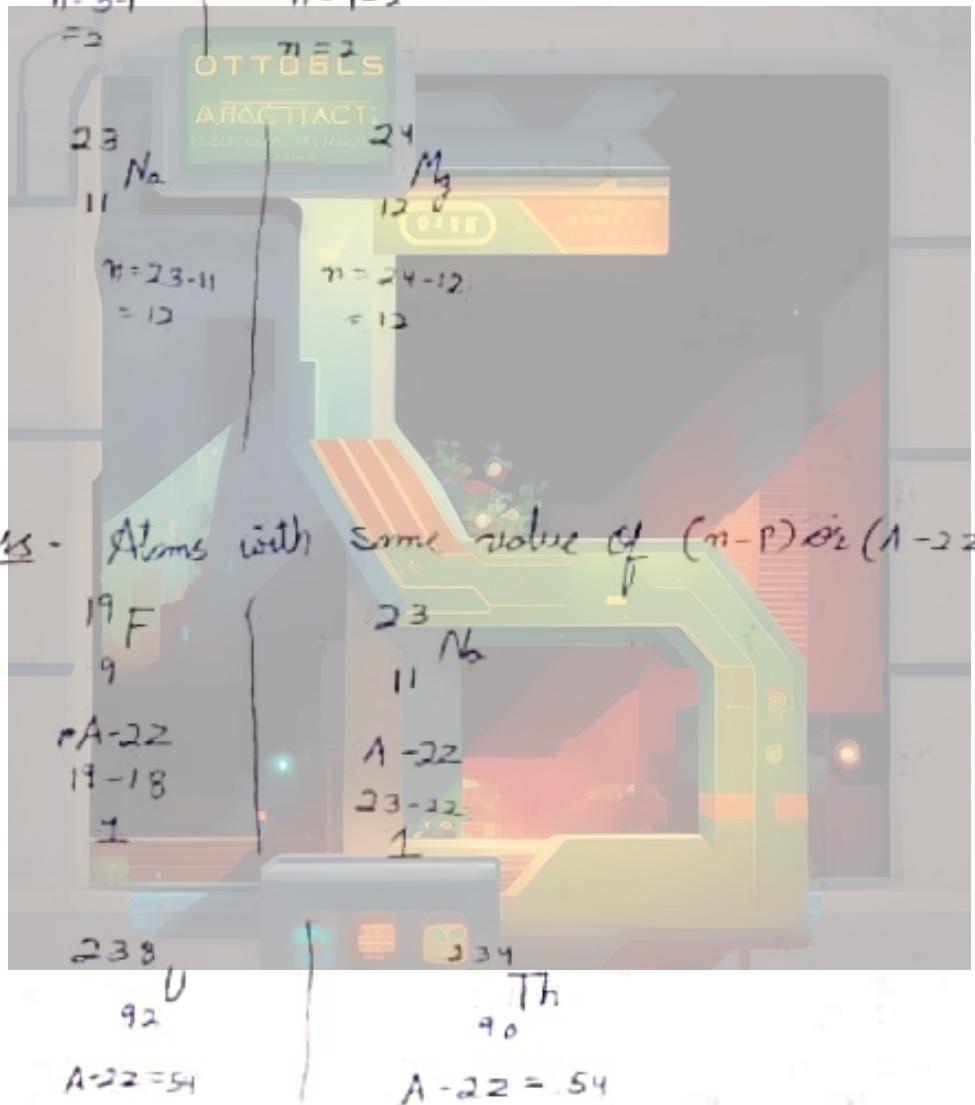
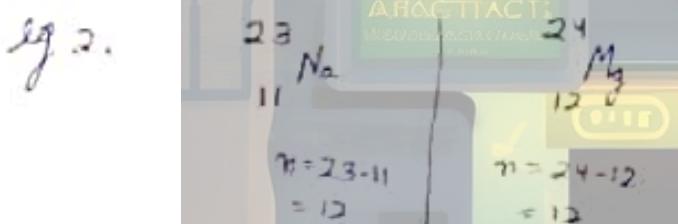
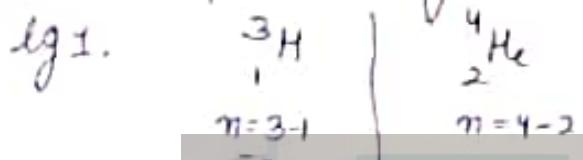
1. Isotopes - Atoms having same atomic number (z) of some element but different mass numbers (A).



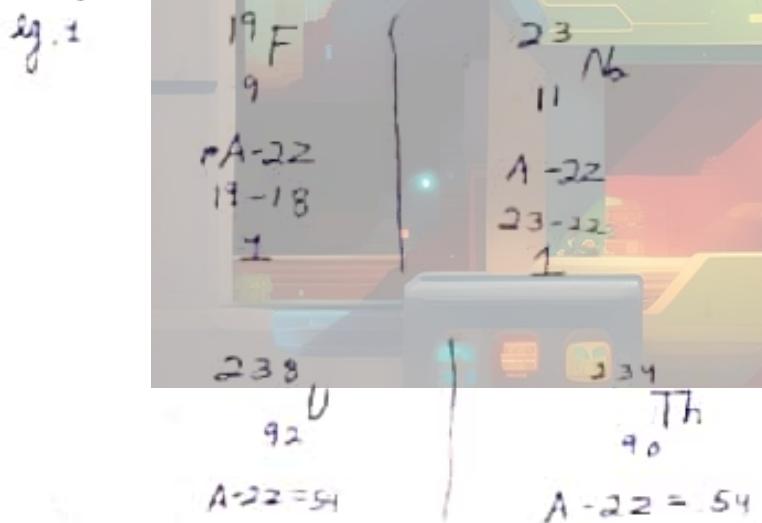
2. Isobars - Atoms having some mass number but different atomic number, atoms of different elements.



3. Isotones: Species having some number of neutrons.

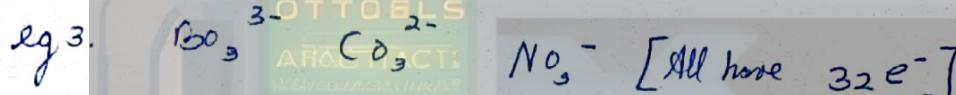
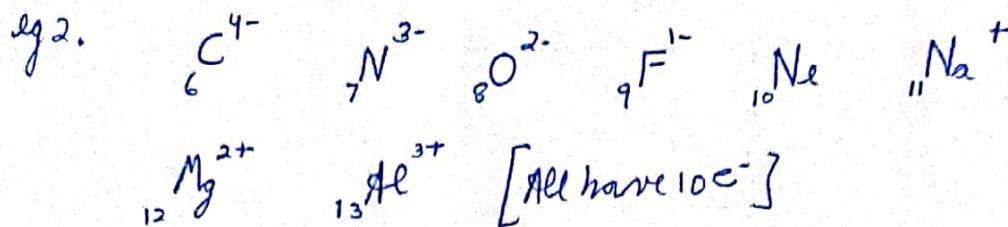
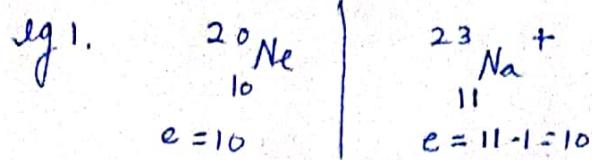


4. Isodisotopes - Atoms with same value of $(n-p)/2$ ($1-22$)



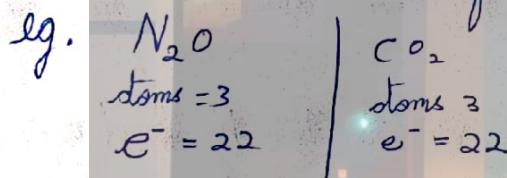
→ atom/ion/molecule

5. Iso-Electronic - Species having same number of total electrons.



* All iso-isotopes are iso-electronic.

6. Isoesters - Species having some number of atoms as well as same no. of electrons.



* All Isoesters are Iso-electronic but not all Iso-electronics are Isoesters.

H.W.

Ez. O-1 (Q1-Q9)

O-2 (Q1)

ES-2 (Q6, 7)

JM (Q1, 6, 7, 9, 11)

O-1

(Q1. B)

Q2 C)

Q3. C)

Q4. B)

O-2

(Q1. Q2 Q3 d)

S-2

(Q6-C)

Q7-D)

JEE-Mains

(Q1. Y)

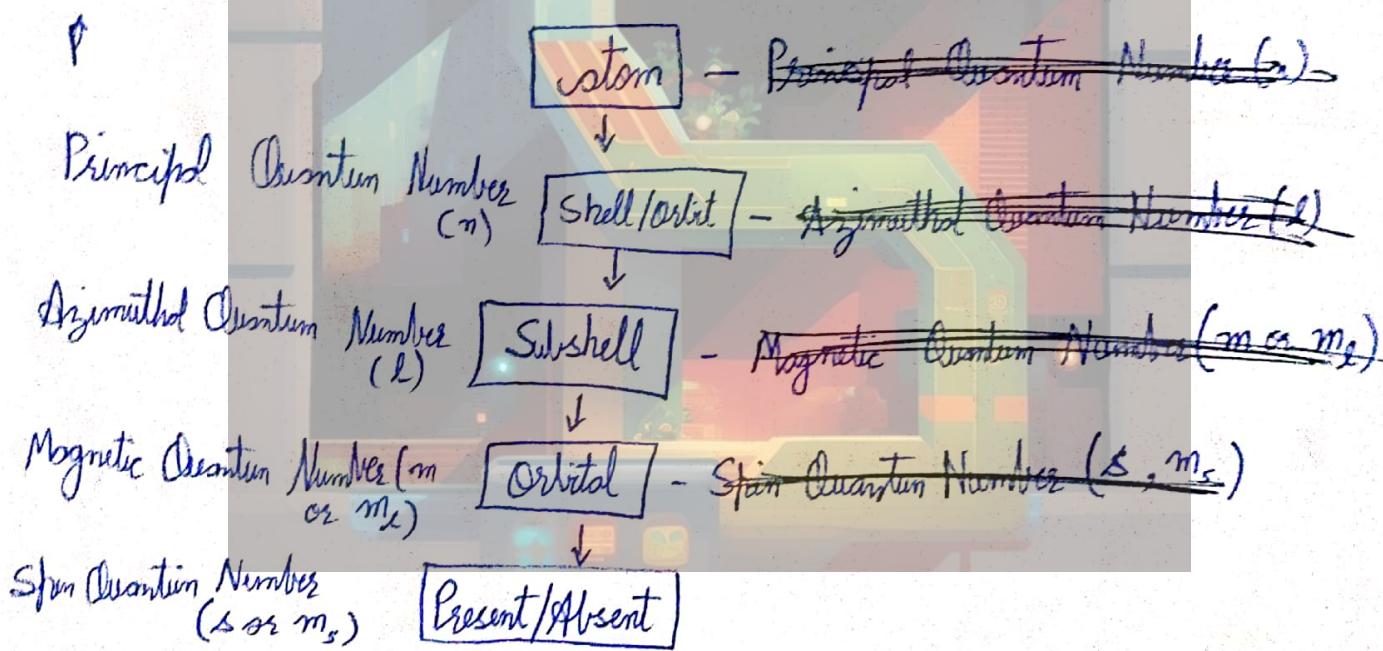
(Q26. I)

(Q27. 4)

(Q9. 4)

(Q11. 2)

Quantum Numbers



* First 3 Quantum Numbers gives information about the orbitals

* First 4 Quantum Numbers gives information about the electrons

(B)

Orbitals - It is defined as 3D space occupied by the electrons around the nucleus where the probability of finding the electron is maximum.

1. Principal Quantum Number (n) - Principal Quantum Numbers represent Shell/orbit/ energy level in which electron is present and its distance from the nucleus. (size of shell/orbit)

* It is also called as Primary Quantum number.

* Permissible values of n - $0, 1, 2, \dots, \infty$

$\downarrow \quad \downarrow \quad \downarrow \quad \downarrow \quad \downarrow \quad \downarrow \quad \dots \quad \downarrow$
K L M N O P

* Maximum number of possible shells in an atom is infinite.

$n \downarrow$ Attraction by Nucleus \uparrow Stability \uparrow Energy \downarrow

* n decides the energy of an electron

* As n increases, energy also increases

* Principal quantum number also gives the idea about Total number of orbitals & total no. of electrons in a shell.

* In a electron orbital, There can be maximum $2 e^-$

* Total no. of orbitals in a shell can be given by n^2 .

* Total no. of electrons in a shell are $2 n^2$

Shell no.	Name in of orbitals	No. no. of electrons
$n=1 (K)$	$\approx 1^2$ ≈ 1	$\approx 2(1)^2$ ≈ 2
$n=2 (L)$	$\approx 2^2$ ≈ 4	$\approx 2(2)^2$ ≈ 8
$n=3 (M)$	$\approx 3^2$ ≈ 9	$\approx 2(3)^2$ ≈ 18
$n=4 (N)$	$\approx 4^2$ ≈ 16	$\approx 2(4)^2$ ≈ 32
$n=5 (O)$		$\approx 2(5)^2$ ≈ 50

2. Azimuthal Quantum Number (l) / secondary Q. No. / subshells Q. No. -

1. The value of $l \Rightarrow 0$ to $(n-1)$ in integral step
2. This Quantum Number gives the information about subshell

- a) name of subshell
- b) energy of subshell (lT , e_n)
- c) shape of subshell

Value of l	Name of Subshell	Shape of subshell
$l=0 (s)$	sharp	spherical
$l=1 (p)$	Principal	Dumbbell
$l=2 (d)$	Diffused	Double Dumbbell
$l=3 (f)$	fundamental	Complex (not in IEE)
$l=4 (g)$	-	-

* The number of subshells in n^{th} shell = n .

n	l (0 to $n-1$)	Subshells
$n=1$	$l=0$ (s)	1s (1 subshell)
$n=2$	$l=0$ (s) 1 (p)	2s 2p (2 subshells)
$n=3$	$l=0$ (s) 1 (p) 2 (d)	3s 3p 3d (3 subshells)

* For existence of a subshell, ($n > l$)

2s - exist

3p - exist

2d - not exist

4g - not exist

* For a particular energy level, the energy of subshell is in following order -

$$[n\ s < n\ p < n\ d < n\ f]$$

$$\text{eg. } 4s < 4p < 4d < 4f$$

3. Magnetic Quantum Number (m or m_l)

1. It gives the information about possible orientation (arrangement) of an ~~electron~~ orbit in space.

2. Values of m depends upon the value of l and ranges from $-l$ to $+l$ (including 0) in integral steps.

O-1 (15, 5, 7)

O-2 (2, 5)

JEE-M (5, 8, 13, 14, 15, 10)

JEE-A (1, 2)

Answers

O-1

(Q5, B)

(Q7, C)

(Q15, D)

O-2

(Q2, C)

(Q5, B) D)

Q J-M

(Q5, I)

(Q8, 2)

(Q10, 2)

(Q13, A)

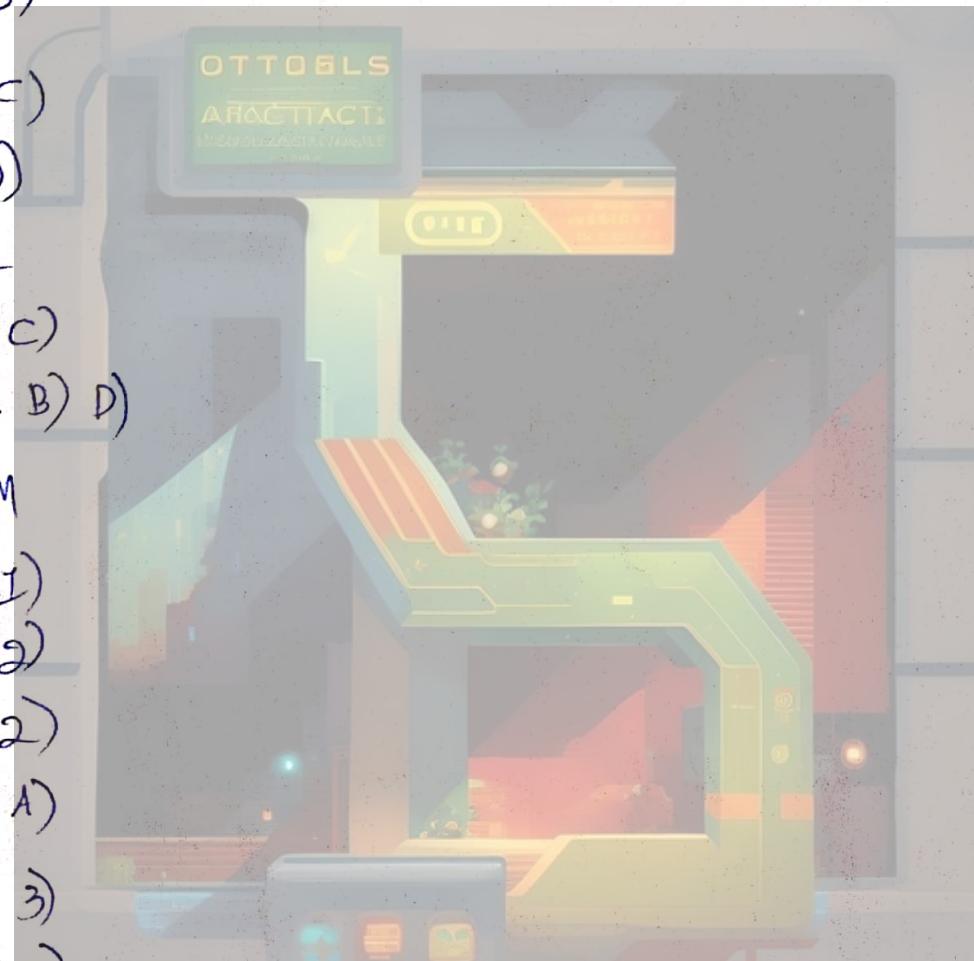
(Q14, 3)

(Q15, C)

J-A

(Q1, 9)

(Q2, 6)



n	l	m	no. of orbitals	no. of e^-
$n=1$	$l=0(s)$	$m=0$	$1=(1)^2$	$1 \times 2 = 2$
$n=2$	$l=1(p)$	$m=-1, 0, +1$	$(2)^2 = 4$	$4 \times 2 = 8$
$n=3$	$l=2(d)$	$m=-2, -1, 0, 1, 2$	$(3)^2 = 9$	$9 \times 2 = 18$
$n=4$	$l=3(f)$	$m=-3, -2, -1, 0, 1, 2, 3$	$(4)^2 = 16$	$16 \times 2 = 32$

* For any given subshell, maximum number of orbitals are $(2l+1)$ & max no. of e^- are $2(2l+1)$

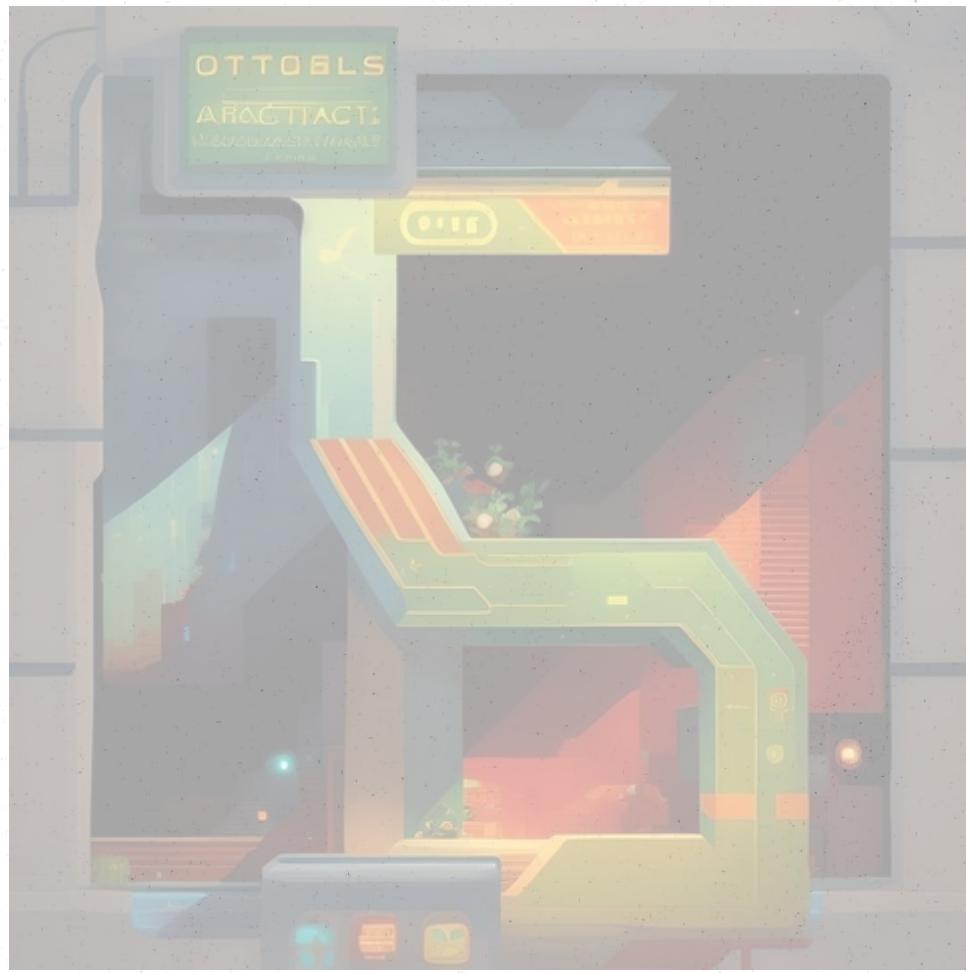
e.g.

$l=0(s)$
 $m=0$
 $2 \times 0 + 1 = 1$ orbital
 1 orbital

$l=1(p)$
 $2 \times 1 + 1 = 3$ orbitals

l	m	representation of orbital
$l=0(s)$	$m=0$	
$l=1(p)$	$m=-1, 0, +1$	
$l=2(d)$	$m=-2, -1, 0, 1, 2$	









Q. If an orbital has $m = -1$ Then it cannot be present in which subshell

- a) s b) p c) d d) f

4. Spin Magnetic Quantum Number (σ, m_s)

1. It is defined to explain the rotation or movement of electron around its own axis.

Clockwise	\uparrow or \downarrow	Anticlockwise
Spin		\downarrow or \uparrow

Spin down	\downarrow or \uparrow	\uparrow or \downarrow
-----------	----------------------------	----------------------------

$\uparrow\downarrow \Rightarrow$ Parallel Spin

$\uparrow\uparrow$ or $\downarrow\downarrow \Rightarrow$ Opposite Spin

Summary of Quantum Numbers

1. $n \rightarrow$ Shell $\rightarrow 1 - \infty$
2. $l \rightarrow$ Subshell $\rightarrow 0$ to $n-1$
3. $m_l \rightarrow$ Orbital $\rightarrow -l$ to $+l$ (including zero)
4. $m_s \rightarrow$ Spin of $\sigma \rightarrow \frac{1}{2}, -\frac{1}{2}$

Q1. What is the highest value of l for shell no. 3? 2

$$\begin{array}{c} n=3 \\ l=0, 1, 2 \\ l=0, 1, 2 \\ m=-1, 0, +1 \end{array}$$

Q2. What is the highest value of m for P subshell? $\square +1$

Q3. What is the minimum value of m for d-subshell? $\square -3$

$$m = -2, -1, 0, 1, 2$$

Q4. Max no. of e^- in S-subshell

$$l=0$$

$$e^- = 6 \quad (2l+1)^2$$

$$= 6(1)^2$$

$$\boxed{l=2}$$

Q5. Max no. of e^- in P-orbitals?

$$l=1$$

$$e^- = (2l+1)^2$$

$$\boxed{l=3}$$

Q6. Select correct set of Quantum Nos.

① n l

	n	l	m	s
(A)	3	3	-2	$-\frac{1}{2}$
(B)	2	1	-2	$+\frac{1}{2}$
✓(C)	3	2	1	$+\frac{1}{2}$
(D)	4	2	-2	$+\frac{1}{2}$

Q7. How many nos. of e^- can be described by Q. Nos.

$n=5, l=2$ in a atom.

$$l=2$$

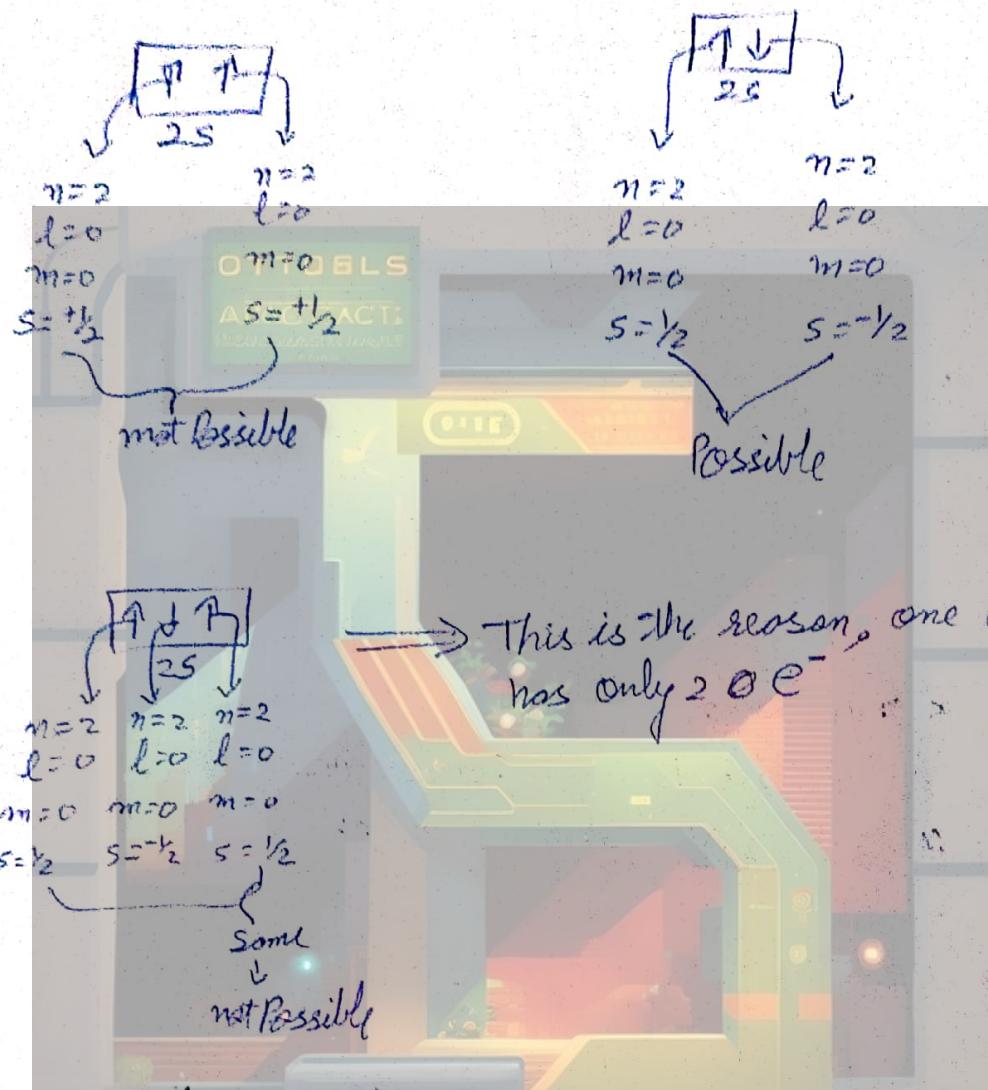
$$m = -2, -1, 0, 1, 2$$

$$\text{no. of } e^- = 2(2l+1)$$

$$\boxed{l=10}$$

Rule for filling electrons

1. Pauli's exclusion Rule - According to this rule, no two electrons in an atom can have all four Quantum Numbers same, at least 1 Quantum Number should be different.



2. Affbau's Rule :- (only for multi electronic species)

→ Not for single electronic species like (H , He^+ , Li^{2+} , Be^{3+})

→ These rules will tell us about the filling of electrons in different subshells,

→ According to this rule, electrons are filled in lower energy subshell first and once they are filled, then filling of next higher energy subshell occurs.

$(n+l)$ rule \rightarrow Higher the value of $(n+l)$ for a subshell, higher will be its energy

Ex - 1

4S
 $n=4$
 $l=0$
 $n+l=4$

3d
 $n=3$
 $l=2$
 $n+l=5$

Ex - 2
5P
 $n=5$
 $l=1$
 $n+l=6$

4f
 $n=4$
 $l=3$
 $n+l=7$

$3d > 4s$ (energy)

$4f > 5P$ (energy)

Ex - 3
4P
 $n=4$
 $l=1$
 $n+l=5$

5S
 $n=5$
 $l=0$
 $n+l=5$

$5S > 4P$ (energy)

* If two subshells having some value of $(n+l)$, then subshell with higher value of n has higher energy.

Q energy order?

1. 3S $n=3$ $l=0$ $n+l=3$	2P $n=2$ $l=1$ $n+l=3$	3d $n=3$ $l=2$ $n+l=5$	4f $n=4$ $l=3$ $n+l=7$
------------------------------------	---------------------------------	---------------------------------	---------------------------------

$4f > 3d > 3s > 2P$

* for calculating no. of e^- in n^{th} period

$ns < (n-2) f < (n-1) d < nP$

ii)

 $4f$
87 $5d$
87 $6s$
6 $4p$
5

$$\boxed{5d > 4f > 6s > 4p}$$

iii)

$$1s < 2s < 3s < 3p < 3d < 4s < 4p < 4d < 4f$$

$$< 5s < 5p < 5d < 5f$$

$$5f > 5d > 4f > 5p > 4p > 5s > 4s > 3p > 3s > 2p > 2s$$
 $> 1s$ n_{shell}

0

1

2

3

4

1

 $1s$

2

 $2s \quad 2p$

3

 $3s \quad 3p$

3d

4

 $4s \quad 4p$

4d

 $4f$

5

 $5s \quad 5p$

5d

 $5f$ $5g$

6

 $6s \quad 6p$

6d

 $6f$ $6g$

7

 $7s \quad 7p$

Q

(23)

Q1. $4d_{xy}$ $4f$ $3d_{xyz}$ $3d_{z^2}$ $5p_y$
 6 7 5 5 6

$4f > \cancel{4d_{xy}}$ $5p_y > 4d_{xy} > \boxed{3d_{yz} = 3d_{z^2}}$,
 degenerate orbitals

degenerate orbitals - All the orbitals of a particular subshell have same energy.

- The orbitals having some value of n & l but different value of m_l have some energy in the absence of an electric & magnetic field.
- These orbitals having same energy of a particular subshell are known as ~~one~~ degenerate orbitals.

Degeneracy :- Total no. of orbitals with same energy.

Q. find degeneracy

- p subshell - 3
- d subshell - 5
- f subshell - 7
- s subshell - ~~10~~ Not Defined ($D=0$)

* For single electron species, or hydrogen like species,
 The energy depends only on n .

H. W.

09-4-24

O-1 (Q5, 7, 15)

O-2 (Q2, Q5)

~~Q8~~ JM (Q5, 8, 10, 13, 14, 15)

JA (Q1, 2)

Answers:-

O-1

~~Q5~~ Q5 - B)

(Q7 C)

(Q15. D) OTTOBLS
ARCTIC

O-2

(Q-2 C)

(Q5. B, D)

Q JM

(Q5 1)

(Q8 2)

(Q10 2)

(Q13 A)

(Q14 3)

(Q15 C)

JEE - Advance

Q1. 9

Q2. 6

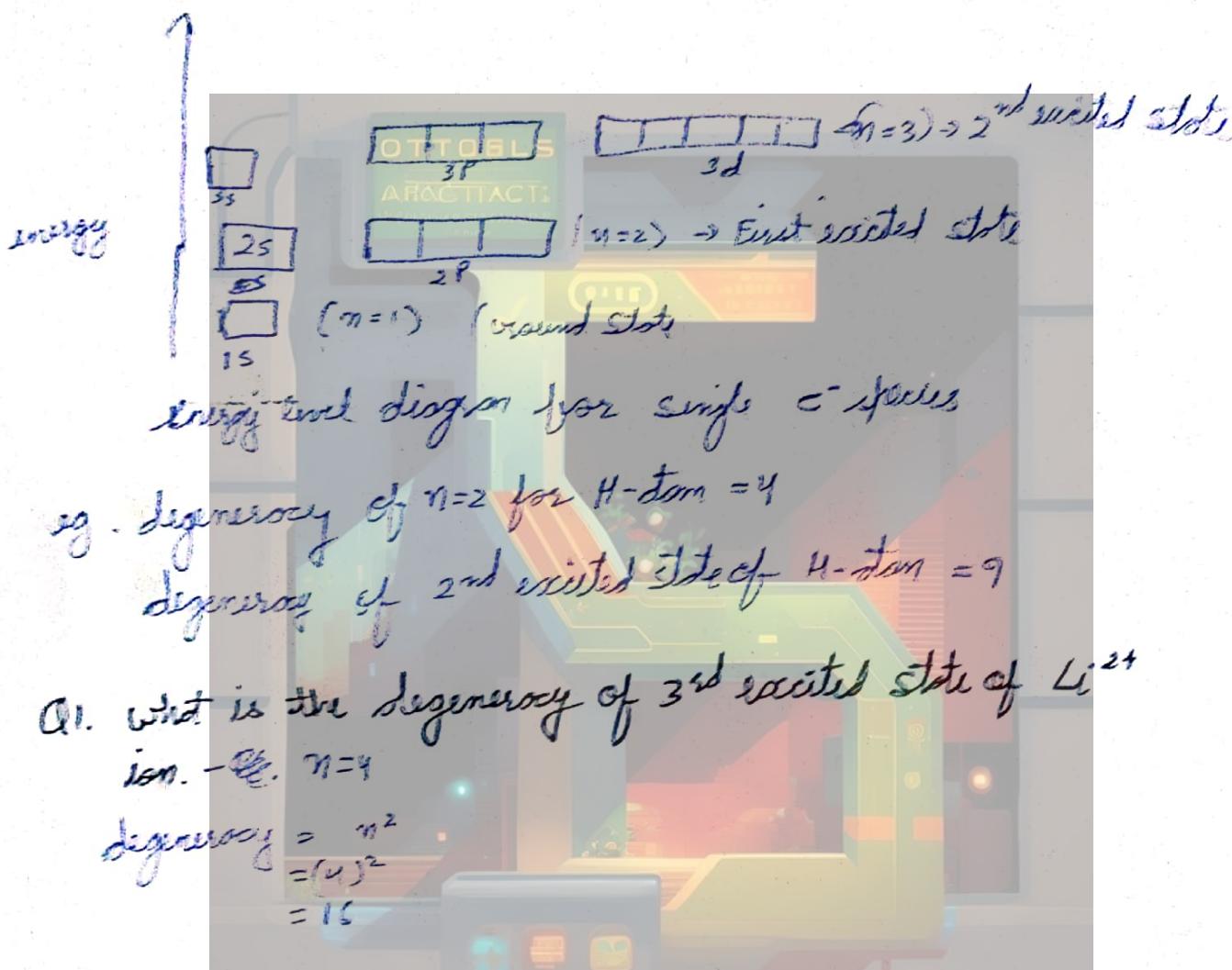
Single Electronic Species (H-like ion species)

o no. of $e^- = 1$

Ex. (H , He^+ , Li^{2+} , Be^{3+} , ...)

→ Energy will depend only on value of n

Energy Order: $1s < (2s=2p) < (3s=3p=3d) < \dots$



Q1. What is the degeneracy of 3rd excited state of Li^{2+} ion. - Ans. $n=4$

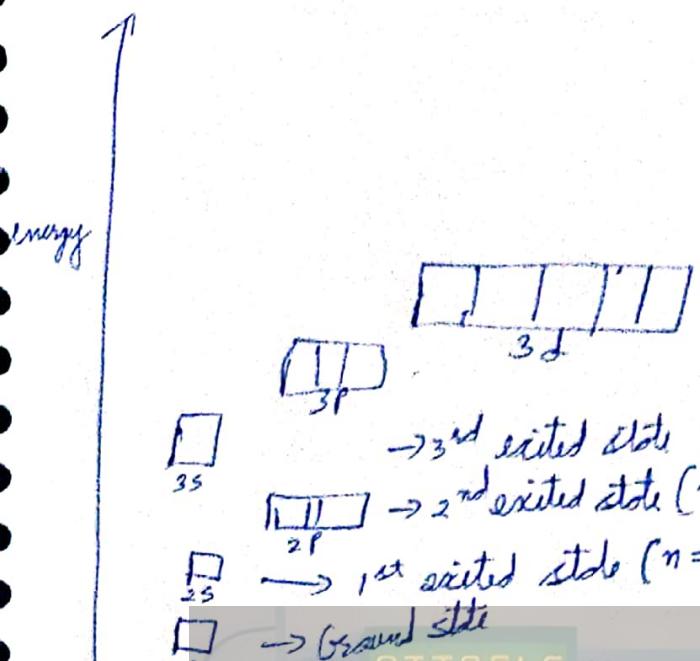
$$\begin{aligned} \text{degeneracy} &= n^2 \\ &= (4)^2 \\ &= 16 \end{aligned}$$

Multi-Electronic Series

No. of $e^- > 1$

eg. He , Ne , Li , Be , ...

Energy of e^- depends on value of n, l, l



energy diagram for multi-electronic species

- Q what is the degeneracy of 2nd excited state of Helium -3
- Q what is the ~~degeneracy~~ of 3rd excited state of Lithium -1

Hund's Rule of maximum multiplicity

- Applies to filling of e^- into degenerate orbitals
- According to this rule electrons are filled into diff orbitals of a subshell $ss \text{ or } s_2$ give maximum no. of e^- unpaired & e^- with parallel spin
- Paring of e^- do not take place in orbitals until each orbital of the same subshell is filled with $1e^-$ of same spin.
- According to this rule, se^- are filled into diff orbitals of a subshell $ss \text{ or } s_2$ give maximum spin multiplicity.

$$\boxed{1} \boxed{1} \boxed{1} \boxed{1} \boxed{1} \quad \checkmark$$

$$\boxed{1} \boxed{1} \boxed{0} \boxed{1} \boxed{1} \quad \checkmark$$

$$\boxed{1} \boxed{1} \boxed{1} \boxed{1} \boxed{1} \quad \times$$

$$\boxed{1} \boxed{1} \boxed{1} \boxed{1} \boxed{1} \quad \checkmark$$

$$\boxed{1} \boxed{1} \boxed{1} \boxed{1} \boxed{1} \quad \times$$

$$\boxed{1} \boxed{1} \boxed{1} \boxed{1} \boxed{1} \quad \checkmark$$

Total Spin (S_T) ↗

$$\boxed{1} \boxed{1} \boxed{1}$$

$$|S_T| = \left| +\frac{1}{2}, +\frac{1}{2}, +\frac{1}{2}, +\frac{1}{2} \right\rangle = \boxed{\frac{3}{2}}$$

$$\boxed{1} \boxed{1} \boxed{1}$$

$$|S_T| = \left| +\frac{1}{2}, -\frac{1}{2}, +\frac{1}{2}, +\frac{1}{2} \right\rangle = \boxed{\frac{1}{2}}$$

Spin Multiplicity (S_m)

$$S_m = 2|S_T| + 1$$

$$\boxed{1} \boxed{1} \boxed{1}$$

$$|S_T| = \frac{3}{2}$$

$$S_m = 2 \times \frac{3}{2} + 1$$

$$\underline{S_m = 4}$$

$$\boxed{1} \boxed{1} \boxed{1} \boxed{1}$$

$$|S_T| = 1$$

$$S_m = 2 \times 1 + 1$$

$$\underline{S_m = 3}$$

Q1. The orbital diagram in which violated Aufbau's principle is

- i) $\begin{array}{|c|} \hline 1L \\ \hline \end{array}$ $\begin{array}{|c|c|} \hline 1L & 1 \\ \hline \end{array}$
- ii) $\begin{array}{|c|} \hline 1 \\ \hline \end{array}$ $\begin{array}{|c|c|c|} \hline 1L & 1 & 1 \\ \hline \end{array}$ ✓
- iii) $\begin{array}{|c|} \hline 1L \\ \hline \end{array}$ $\begin{array}{|c|c|c|} \hline 1 & 1 & 1 \\ \hline \end{array}$
- iv) $\begin{array}{|c|} \hline 1L \\ \hline \end{array}$ $\begin{array}{|c|c|c|} \hline 1 & 1 & 1 \\ \hline \end{array}$

Q2. The orbital diagram in which Hund's Rule is violated is

- i) $\begin{array}{|c|c|c|c|c|} \hline 1 & 1 & 1 & 1 & 1 \\ \hline \end{array}$
- ii) $\begin{array}{|c|c|c|c|} \hline 1 & 1 & 1 & 1 \\ \hline \end{array}$ ✓
- iii) $\begin{array}{|c|c|c|} \hline 1L & 1L & 1 \\ \hline \end{array}$ ✓
- iv) $\begin{array}{|c|c|c|c|c|} \hline 1 & 1 & 1 & 1L & 1 \\ \hline \end{array}$

Q3. Match

Configuration

- i) $\begin{array}{|c|c|} \hline 1L & 1L \\ \hline \end{array}$
- A i) $\begin{array}{|c|} \hline 1L \\ \hline \end{array}$ $\begin{array}{|c|} \hline 1L \\ \hline \end{array}$ $\begin{array}{|c|c|} \hline 1 & 1 \\ \hline \end{array}$
- B ii) $\begin{array}{|c|} \hline 1L \\ \hline \end{array}$ $\begin{array}{|c|} \hline 1L \\ \hline \end{array}$ $\begin{array}{|c|c|} \hline 1 & 1 \\ \hline \end{array}$
- C iii) $\begin{array}{|c|} \hline 1 \\ \hline \end{array}$ $\begin{array}{|c|} \hline 1L \\ \hline \end{array}$ $\begin{array}{|c|c|c|} \hline 1 & 1 & 1 \\ \hline \end{array}$
- D iv) $\begin{array}{|c|c|} \hline 1 & 1 \\ \hline \end{array}$ $\begin{array}{|c|} \hline 1 \\ \hline \end{array}$ $\begin{array}{|c|c|c|} \hline 1 & 1 & 1 \\ \hline \end{array}$

Violated Rule

- P) Pauli's Rule
- Q) No Rule violated
- R) Aufbau's Rule
- S) Hund's Rule

- A) Q
- B) P S
- C) R S
- D) P R S

Electronic Configuration of a atom.

→ The distribution of electrons in various shells, subshells and orbitals in an atom is called its electronic configuration.

e.g. $1s^2 \rightarrow$ no. of electrons
 $n=1$ $l=s$
 $=0$

$$\begin{aligned} S &\rightarrow 2 e^- \\ P &\rightarrow 6 e^- \\ d &\rightarrow 10 e^- \\ f &\rightarrow 14 e^- \end{aligned}$$

e.g.

- ${}_1 H \rightarrow 1s^1$
- ${}_2 He \rightarrow 1s^2$
- ${}_3 Li \rightarrow 1s^2 2s^1$
- ${}_4 Be \rightarrow 1s^2 2s^2$
- ${}_5 B \rightarrow 1s^2 2s^2 2p^1$
- ${}_6 C \rightarrow 1s^2 2s^2 2p^2$
- ${}_7 N \rightarrow 1s^2 2s^2 2p^3$
- ${}_8 O \rightarrow 1s^2 2s^2 2p^4$
- ${}_9 F \rightarrow 1s^2 2s^2 2p^5$
- ${}_{10} Ne \rightarrow 1s^2 2s^2 2p^6$

orbital notation method

$\text{C} \rightarrow [\text{He}] 1s^2 2s^2$ [odd stages omitted]

$\text{C} \rightarrow [\text{He}] 2s^2 2p^2$ [combined form]

$_{11}\text{Na} \rightarrow [\text{Ne}] 3s^1$

$_{12}\text{Mg} \rightarrow [\text{Ne}] 2s^2$

$_{13}\text{Al} \rightarrow [\text{Ne}] 3s^2 3p^1$

$_{14}\text{Si} \rightarrow [\text{Ne}] 3s^2 3p^2$

$_{15}\text{P} \rightarrow [\text{Ne}] 3s^2 3p^3$

$_{16}\text{S} \rightarrow [\text{Ne}] 3s^2 3p^4$

$_{17}\text{Cl} \rightarrow [\text{Ne}] 3s^2 3p^5$

$_{18}\text{Ar} \rightarrow [\text{Ne}] 3s^2 3p^6$

$_{19}\text{K} \rightarrow [\text{Ar}] 4s^1$

$_{20}\text{Ca} \rightarrow [\text{Ar}] 4s^2$

$_{21}\text{Sc} \rightarrow [\text{Ar}] 4s^2 3d^1$

$_{22}\text{Ti} \rightarrow [\text{Ar}] 4s^2 3d^2$

$_{23}\text{V} \rightarrow [\text{Ar}] 4s^2 3d^3$

$_{24}\text{Cr} \rightarrow \cancel{[\text{Ar}] 4s^2 3d^4}$

$_{25}\text{Mn} \rightarrow [\text{Ar}] 4s^2 3d^5$

$_{26}\text{Fe} \rightarrow [\text{Ar}] 4s^2 3d^6$

$_{27}\text{Co} \rightarrow [\text{Ar}] 4s^2 3d^7$

$_{28}\text{Ni} \rightarrow [\text{Ar}] 4s^2 3d^8$

$_{29}\text{Cu} \rightarrow \cancel{[\text{Ar}] 4s^2 3d^9}$

$_{30}\text{Zn} \rightarrow [\text{Ar}] 4s^2 3d^10$



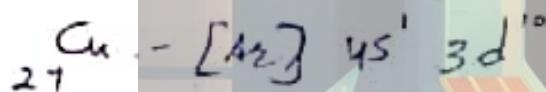
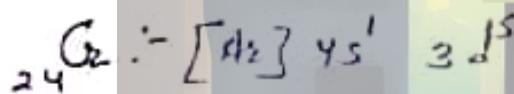
Frost Gas.

He
Ne
Ar
Kr
Xe
Rn

Atomic no.

2
10
18
36
54
86

Exception



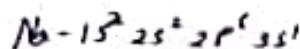
Valence Shell (n)

- * It is the outermost shell which is occupied by a electron
- * The highest value of n in the electronic configuration is valence shell.

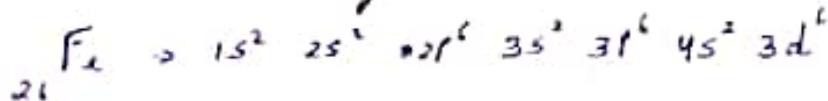
Valence Shell / outermost shell (n) = 3rd

Penultimate Shell ($n-1$) = 2nd

Pre-Penultimate Shell ($n-2$) = 1st



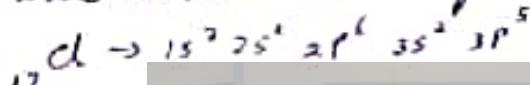
Q1. write down the all four quantum No. for last e⁻ of Fe (atom)



$$\checkmark \delta^- \rightarrow n=4 \\ l=0 \\ m=0 \\ s=\pm\frac{1}{2} / \mp\frac{1}{2}$$

$$\text{last } e^- \rightarrow n=3 \\ l=2 \\ m=-2, -1, 0, 1, 2 \\ s=\pm\frac{1}{2} / \mp\frac{1}{2}$$

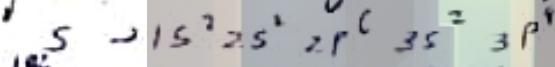
Q2. write write down the all four Q. No. for 7^{th} e⁻ of Cl (atom)



$$\checkmark \delta^- \rightarrow n=3 \\ l=1 \\ m=-1, 0, 1 \\ s=\pm\frac{1}{2} / \mp\frac{1}{2}$$

OTTO BLS $n=2$
ARACTACTL $l=1$
MOCOOGAATK $m=-1, 0, +1$
 $s=\pm\frac{1}{2}$

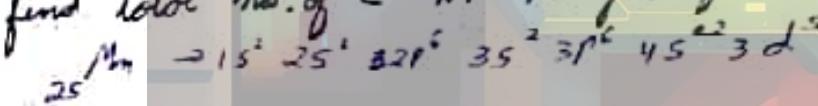
Q3. find ^{maximum} no. of e⁻ in subshell for which $m=0$?



for $m=0$, $l>0$ ~~any~~ l in every subshell.

[10]

Q4. find total no. of e⁻ in monogress for which



a) value of $n=3$
 $= 6 + 5 + 2$
 $\boxed{13} \checkmark$

b) value of $(n+1)=4$
 $= 6 + 2 +$
 $\boxed{8} \checkmark$

b) value of $l=2$
 $= 6 + 6$
 $\boxed{12} \checkmark$

c) value of $(l+m)=0$
 $= 2 + 2 + 2 + 2 + 2 + 2 + 2$
 $\boxed{14} = \boxed{13} \checkmark$

c) value of $|m|=1$
 $= 2 + 2 + 2$
 $= 4 + 4 + 4$
 $\boxed{12} \checkmark$

d) value of $(l+m)=2$
 $= 2 + 2 + 2$
 $\boxed{6} = \boxed{5} \checkmark$

d)

e) $(l \times l \times m)=0$
 $= 8 \times 2 \times 0$
 $\boxed{16} \checkmark$

(33)

& find out minimum no. of e⁻ in magnet for $s = -\frac{1}{2}$

$$10+2 \rightarrow 12 - \boxed{4\bar{5}6}$$

$$\text{min no. of } e^- = 10$$

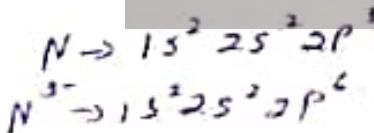
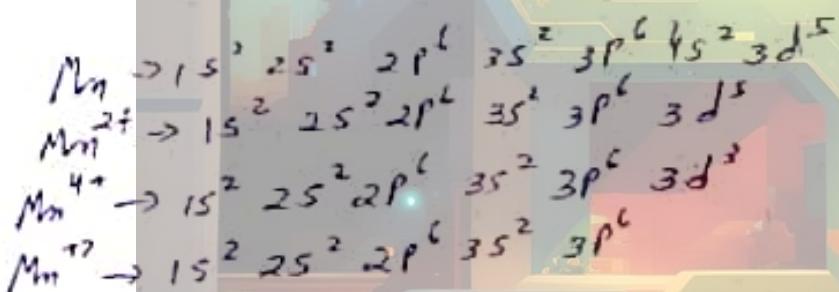
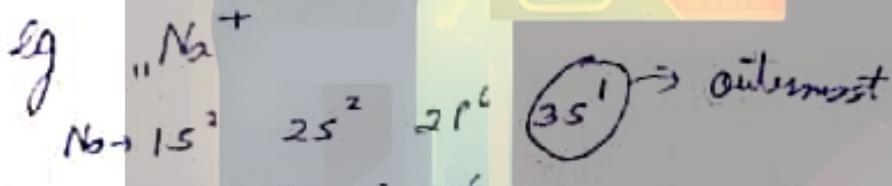
$$\text{max no. of } e^- = 15$$

Electronic configuration of ions

step-1 - first write down the electronic configuration of neutral atom.

step-2 removal of electron takes place from outermost shell
i.e. n^{th} shell.

within a shell removal of electron takes place from
inner subshells which have more value of l (azimuthal
(quantum number))



Extra stability of half filled & fully filled subshell.

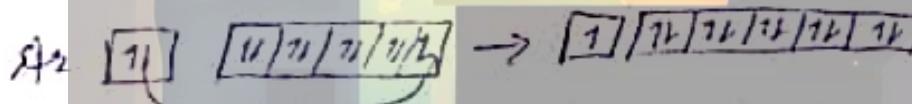
I. The electronic configuration of most of the atom follows Aufbau's rule.

Exception of Aufbau's Rule:-

1. $_{24}^{40}\text{Cr} \rightarrow [Ar] 4s^2 3d^4 \rightarrow \text{expected}$
 $[Ar] 4s^1 3d^5 \rightarrow \text{reality}$



2. $_{29}^{41}\text{Cu} \rightarrow [Ar] 4s^2 3d^9 \rightarrow \text{expected}$
 $[Ar] 4s^1 3d^{10} \rightarrow \text{reality}$



2. Here the two subshells ($4s$ & $3d$) differ slightly in their energy. ($4s < 3d$). An electron shifts from a subshell of lower energy ($4s$) to a subshell of higher energy ($3d$). Provided such a shift results in all orbitals of a subshell of higher energy getting either completely filled or half filled.

Half filled subshell

S - $\boxed{1}$

P - $\boxed{111}$

d - $\boxed{111111}$

f - $\boxed{11111111111111}$

Fully filled subshell

$\boxed{11}$

$\boxed{111111}$

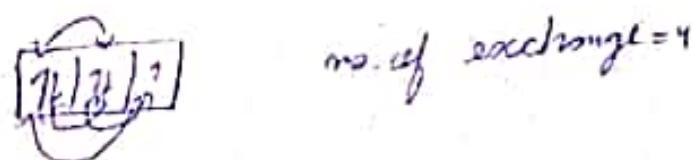
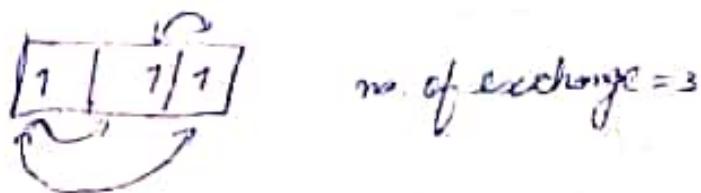
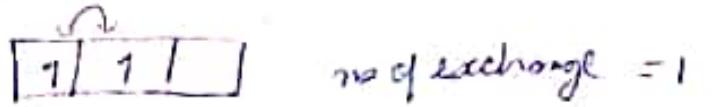
$\boxed{1111111111}$

$\boxed{11111111111111}$

- Note:-
1. Half filled & fully filled e^- configuration is more stable than partially filled e^- configuration.
 2. fully filled e^- configuration is more stable than half filled e^- configuration.
 3. It has been found that there is an extra stability associated with this electronic configuration. This stabilization is due to following 2 factors

Explanation:- Symmetry leads to stability (Nature's law)

Expt 2 - Exchange energy - The electron present in degenerate orbitals having parallel spin can exchange their positions and energy released due to this exchange is known as exchange energy.



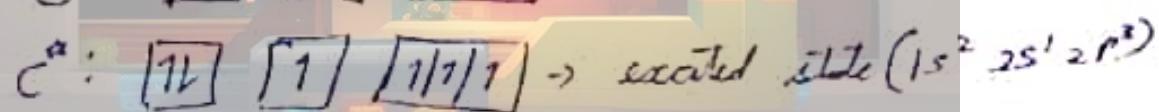
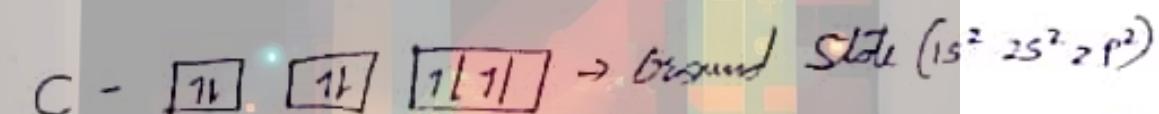
$$\text{no. of exchange} = n(n-1) + \frac{n'(n'-1)}{2}$$

n = no. of e⁻ in up spin
 n' = no. of e⁻ in down spin

No. of exchange $\uparrow \rightarrow$ energy $\downarrow \rightarrow$ stability \uparrow

~~* Excited energy~~ - The minimum amount of energy required to excite an electron from ground state (lower energy level) of an atom to any excited state (higher energy) is called excited energy.

e.g.



Exceptional Electronic Configurations

S. no.	Element	Z	Config
1.	Cr	24	[Ar] 4s ¹ 3d ⁵
2.	Cu	29	[Ar] 4s ¹ 3d ¹⁰
3.	Nb	41	[Kr] 5s ¹ 4d ¹
4.	Mo	42	[Kr] 5s ² 4d ⁵
5.	Pu	94	[Kr] 5s ¹ 4d ⁷
6.	Rh	45	[Kr] 5s ¹ 4d ⁷
7.	Pd	46	[Kr] 4d ¹⁰
8.	Ag	47	[Kr] 5s ¹ 4d ¹⁰
9.	La	57	[Xe] 6s ² 5d ¹
10.	Pt	79	[Xe] 6s ¹ 4f ¹⁴ 5d ⁹
11.	Au	79	[Xe] 6s ¹ 4f ¹⁴ 5d ¹⁰
12.	Hg	89	[Rn] 7s ² 6d ¹
13.	Th	90	[Rn] 7s ² 6d ²

HW 13-04-24 34-Questions

O-1 (Q 6, 8, 9, 10, 11, 13, 16, 17)

O-2 (Q 3, 4, 6, 7, 8, 9, 10, 11, 12, 13, 14, 16)

S-1 (Q 1, 2, 3, 5, 6)

S-2 (Q 1, 2, 3, 9)

JM (Q 2, 4, 12, 16)

JA (Q 3)

JA Q 3 - 3

JM Q 2 - 1)

Q 4 - 3)

Q 72 - 3)

Q 16 - A)

S-2

Q 1 - C

Q 2 - B

Q 3 - C

Q 9 - C

S-1

Q 1 - 4

Q 2 - 9

Q 3 - 8

Q 5, 6 (Raut)

Q 6 - 2. (no of odd bits)

O-2

Q 3 - A, B

Q 4 - C, D

Q 6 - B, D

Q 7 - A, B, D

Q 8 - B, D

Q 9 - A, B, D

Q 10 - C, D

Q 11 - A, B, D

Q 12 - C, D

Q 13 - C, D, B

Q 14 - B, C

Q 16 - A, B

O-1

Q 6 - A

Q 8 - B

Q 9 - C

Q 10 - C

Q 11 - B

Q 12 - B

Q 13 - B, C

Q 15 - C, D

Q 17 - A

Q. Find maximum no. of e^- in s^m shell.

$$m = 5$$

ss , $(s-3)g$, $(s-2)f$, $(s-1)d$, sp

\downarrow \downarrow \downarrow \downarrow \downarrow
2 10 6

$$10 + 6 + 2 \rightarrow 18 \text{ elements}$$

Magnetic Properties

Diamagnetic

Ammagnetic

Spin Magnetic moment (μ_s) -

$$\mu_s (mu)$$

→ The magnetic moment due to spinning of e^- is called spin magnetic moment.

→

$$\mu_s = \sqrt{(n)(n+2)}$$

n : no. of unpaired electrons

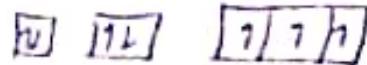
Unit: $B.M.$ (Bohr Magneton)

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

$[2] \quad [2]$ $[2] \quad [2] \quad [2]$

No. of unpaired electrons = $n = 0$

$$\frac{1}{2} - N \rightarrow 1s^2 2s^2 2p^3$$



$$n = 3$$

$$\mu = \sqrt{n(n+2)}$$

$$\mu = \sqrt{15}$$

$$\mu = 3.87$$

no. of unpaired e^- | $\sqrt{n(n+2)}$

μ

$$\underline{0}$$

$$\underline{0}$$

$$\underline{1}$$

$$\underline{1.73}$$

$$\underline{2}$$

$$\underline{2.84}$$

$$\underline{3}$$

$$\underline{3.87}$$

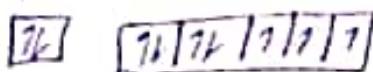
$$\underline{4}$$

$$\underline{4.91}$$

value of μ before decimal (1.73) = no. of unpaired e^-

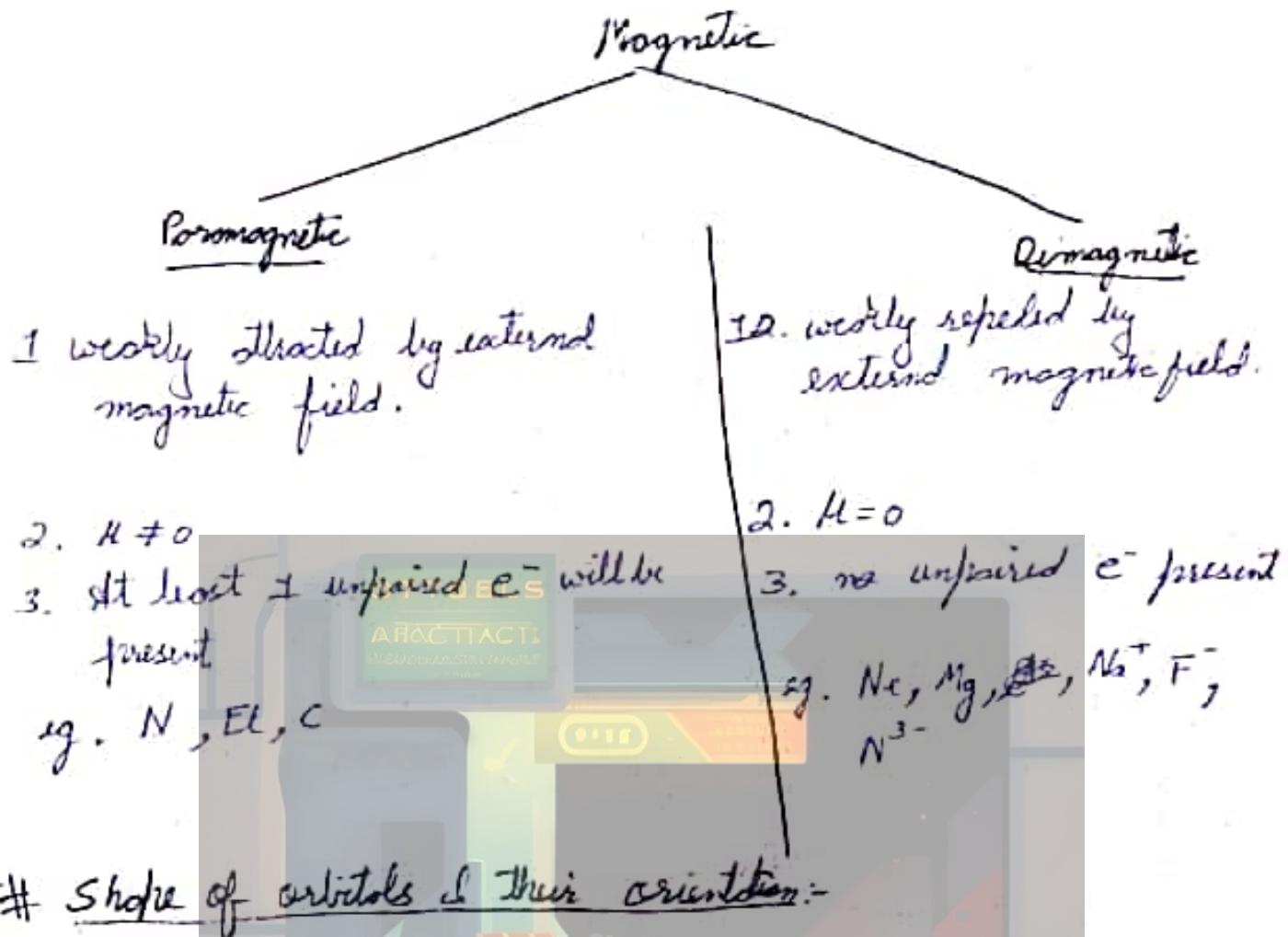
Q If Ca^{2+} has magnetic moment 3.87 B.M. find x .

$$4s^2 3d^x$$



	$+1 \rightarrow$	$[\underline{1} \quad \underline{1} \quad \underline{1} \quad \underline{1} \quad \underline{1}] \rightarrow n = 4$	$x = 1$
	$+2 \rightarrow$	$[\underline{1} \quad \underline{1} \quad \underline{1} \quad \underline{1} \quad \underline{1}] \rightarrow n = 3$	$x = 2$
	$+3 \rightarrow$	$[\underline{1} \quad \underline{1} \quad \underline{1} \quad \underline{\quad}] \rightarrow n = 2$	$x = 6$

$$x = 2, 6$$



Shape of orbitals & Their orientation:-

1. Nodal Plane - It is defined as the plane passing through the nucleus where probability of finding an electron is zero.

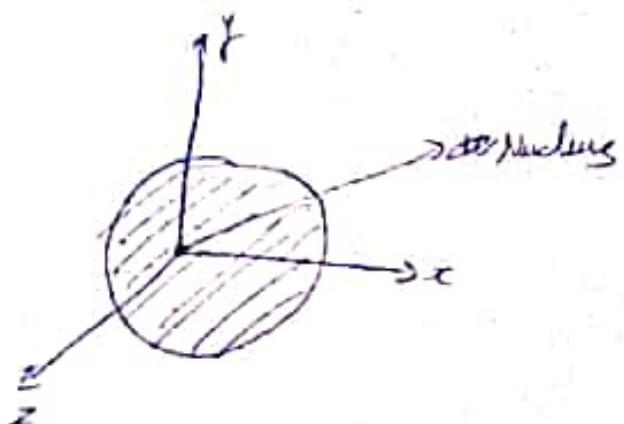
⇒ no. of nodal planes = value of l

S-orbital	$l=0$	N.P. = 0
-----------	-------	----------

P-orbital	$l=1$	N.P. = 1
-----------	-------	----------

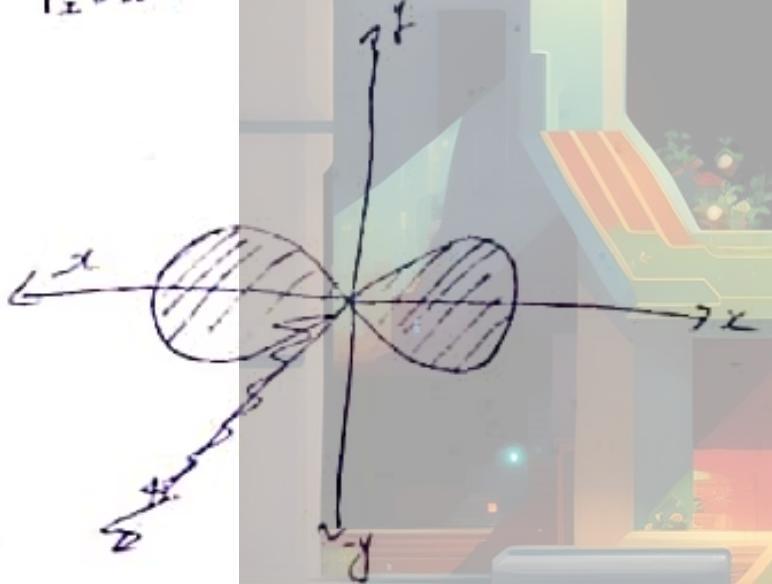
D-orbital	$l=2$	N.P. = 2
-----------	-------	----------

1. S-orbital (spherical)



2. P-orbital (Dumbbell)

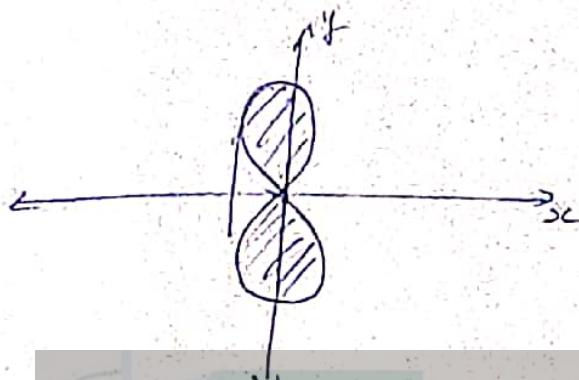
P_z -orbital



$$N.P = 1$$

Nodal plane - αYZ axis or ZY

p_x orbital -



Nodal Plane - XZ or ZX

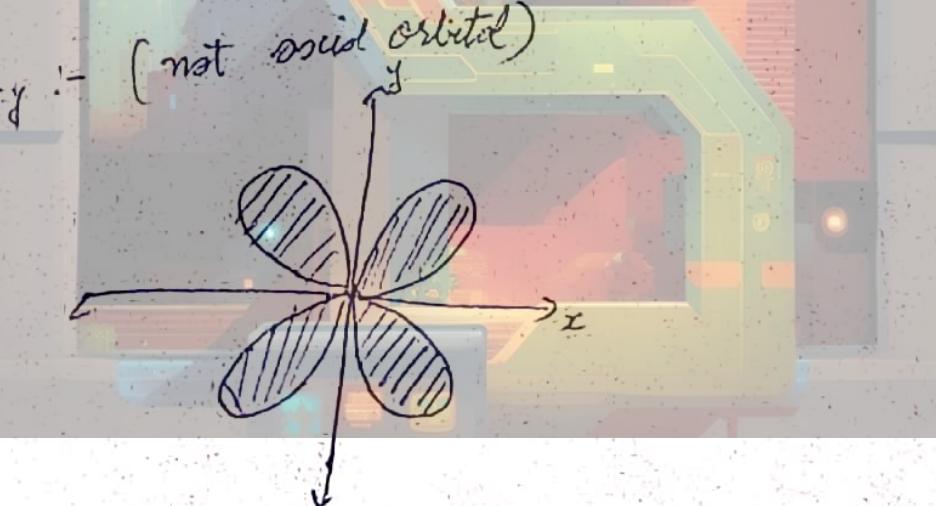
These orbitals are axial orbitals because e⁻ density lies on axes.

p_z orbital:-

Nodal plane - XY or YX

3. d-orbital (double-dumbell) ($d_{xy}, d_{xz}, d_{yz}, d_{z^2-y^2}, d_{x^2-z^2}$)

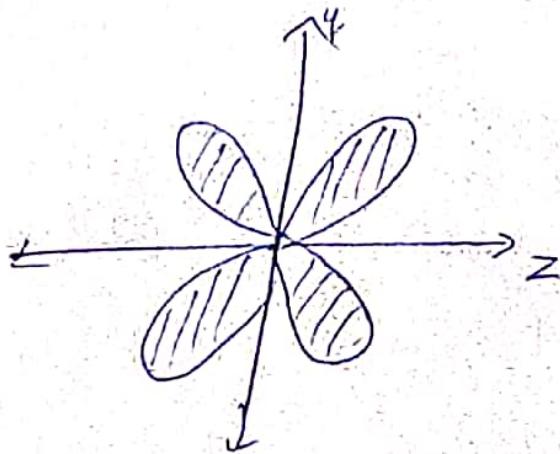
d_{xy} :- (not axial orbital)



Nodal Plane - XZ and YZ

-2

d_{yz} (not axial orbital)



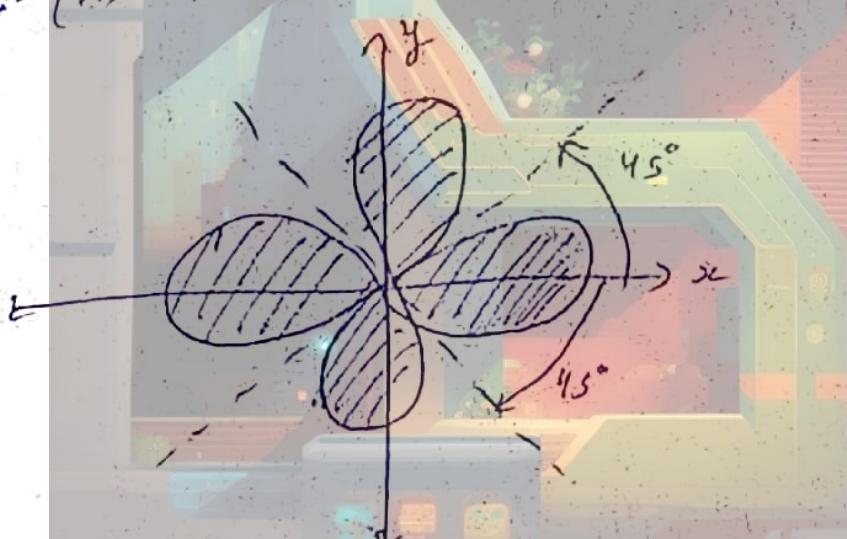
Nodal Planes: - XY, XZ

d_{xz} : - (not axial orbital)

Nodal Planes: - YZ, XY

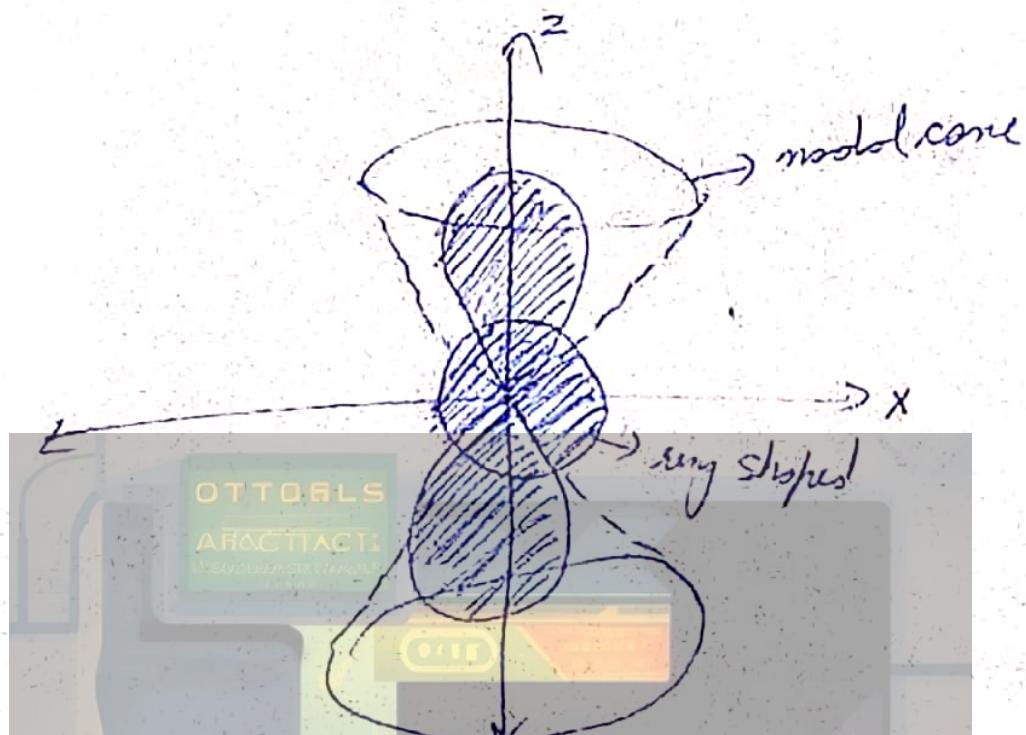
- 2

$d_{x^2-y^2}$: - (Axial orbital)



Nodal Planes: -
- Two nodal plane inclined at 45° from
vertical axis

d_{z^2} :- (Axial orbital)



Nodal Planes :- 0

- No nodal planes

- two nodal cones exist

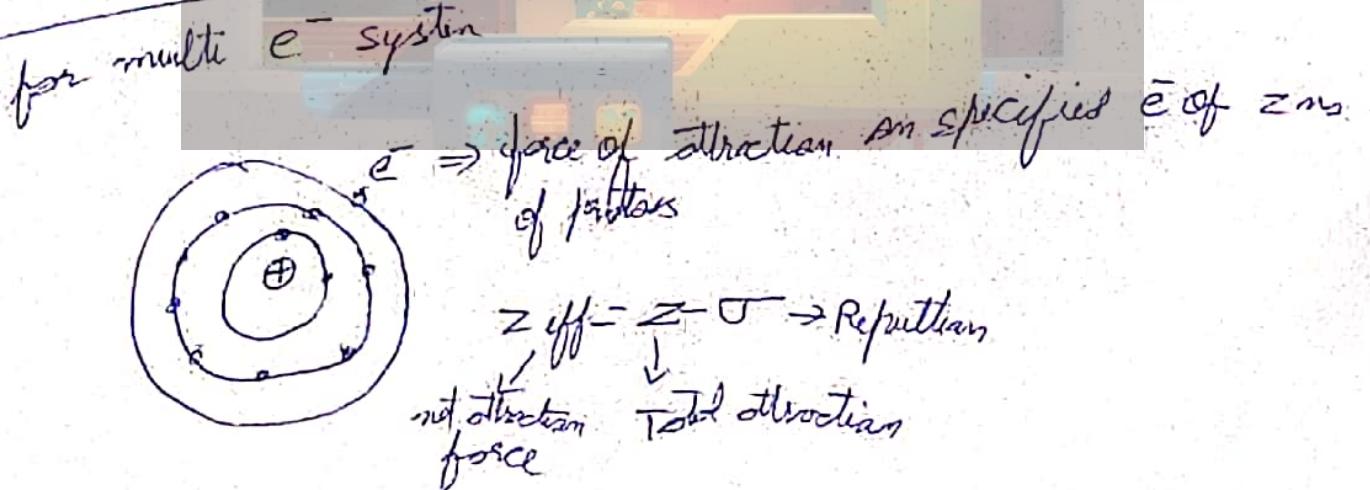
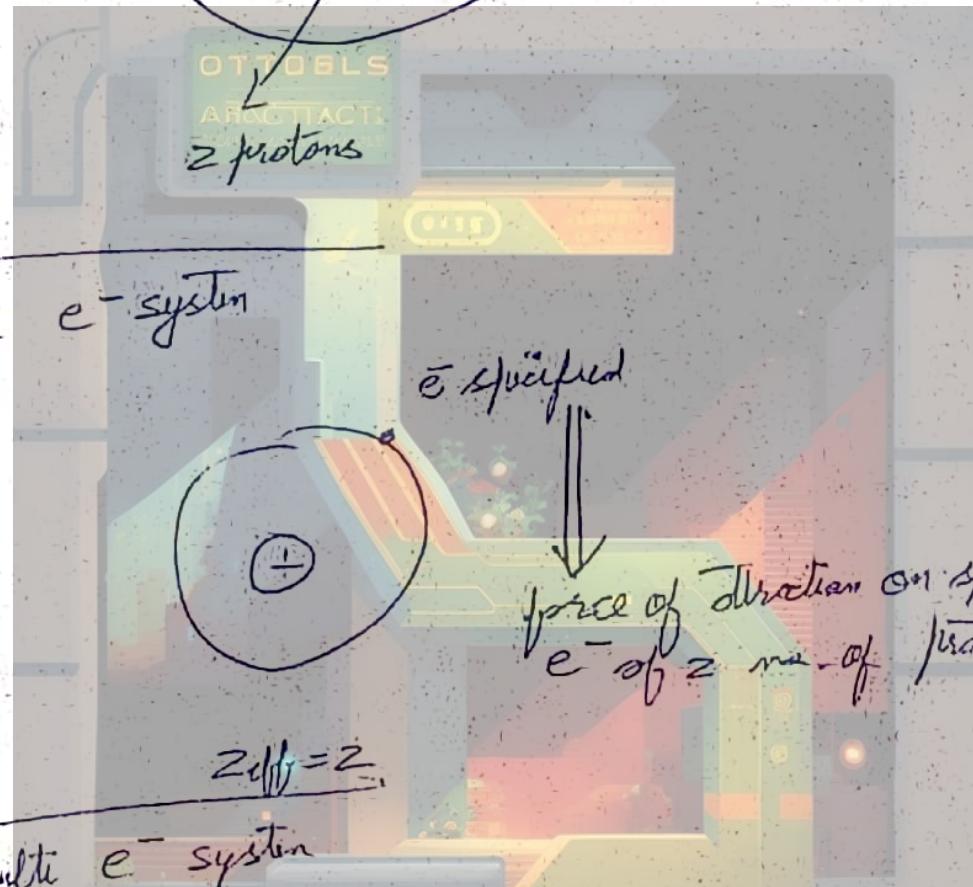
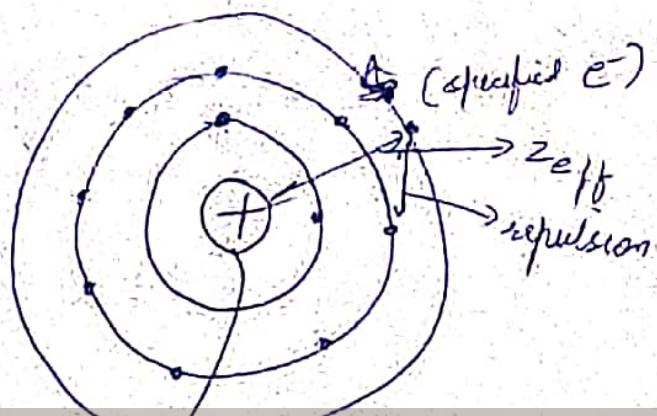
Screening constant & effective nuclear charge

(σ)

↓
sigma

(Z/Z_{eff})

→ (Effective Nuclear charge) -
It is the force of attraction between nucleus and only
specified electron.



$Z_{\text{eff}} = \text{effective Nuclear charge}$

$Z = \text{atomic no.} / \text{no. of proton}$

$\sigma = \text{slater constant} / \text{shielding constant} / \text{screening constant}$

Note:-

→ Due to the presence of inner electrons, the attractive force of nucleus on specified electron is reduced.

→ The inner e^- present in an atom apply a repulsive force on specified e^- . This repulsive force of inner shell e^- is called shielding effect or screening effect.

Calculation of slater constant (σ)

Rule 1 - write electronic config of an atom and arrange e^- in following manner.
↓
electronic config.

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

Rule 2 - Decide the group of e^- for which σ is to be calculated.

2 types of groups

~~n s / n p~~

n s / n p

1. All the e^- present right hand side of the specified e^- contribute nothing to sigma.
2. The remaining e^- in some group ($n s / n p$) contribute 0.35 each to sigma.
3. The e^- present in $(n-1)$ shell contribute 0.85 each to sigma.
4. The e^- present in $(n-2)$ ($n-3$) ... contribute ± 4 each to sigma.

nd / nf

1. e^- present right hand side of No or Nf group contribute nothing to sigma.
2. e^- present in some group (nd / nf) contribute 0.35 each to sigma.
3. All e^- present in LHS of (nd / nf) group contribute \pm each to σ .

Z_{eff} for (1s) e^- of Fe: 3.75
(3d) e^- of Fe: 6.25

It is easier to remove an electron from 4s as compared to 3d because Z_{eff} on 3d is more.

Q1. Calculate value of σ & Z_{eff} for valence shell e^- of Be.

$$\text{Be} - (1s^2)(2s^2) \quad (2s) \rightarrow \text{specified } e^- / \text{lost } e^-$$

\leftarrow

$$\sigma = 0.35 + 2(0.35)$$

$$\sigma = 0.35 + 1.7$$

$$\sigma = \boxed{2.05}$$

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

ARCTIC

$$Z_{\text{eff}} = 4 - 2.05$$

$$\boxed{Z_{\text{eff}} = 1.95}$$

~~Q2~~

Q2. Calculate value of Z_{eff} & σ for $3s e^-$ of Na

$$\text{Na} \rightarrow (1s)^2 (2s^2 2p^6) (3s^1)$$

$$\sigma = (8 \times 0.35)(1 \times 2) + (0 \times 0.35)$$

$$\sigma = 6.80 + 2$$

$$\sigma = 8.8$$

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

$$= 11 - 8.8$$

$$\boxed{= 2.2}$$

(4)

Q3. Calculate value of z_{eff} & σ for 3d e⁻ of iron.
also calculate σ & z_{eff} for 4s e⁻

$$E_e = (1s)^2 (2s, 2p)^{18} (3s, 3p)^{18} (3d)^6 (4s)^2$$

$$3d \sigma = (5 \times 0.35) (1 \times 18)$$

$$= 18 + 1.75$$

$$\boxed{= 19.75}$$

$$3d z_{eff} = 2 - 19.75$$

$$= 26 - 19.75$$

$$\boxed{= 6.25}$$

$$4s \sigma = (1 \times 0.35) (6 \times 0.85) (1 \times 18)$$

$$= 18 + 0.35 + 5.10$$

$$\boxed{= 23.45}$$

$$4s z_{eff} = 26 - 23.45$$

$$\boxed{= 2.65}$$

$$4s \sigma = 0.35 + (14 \times 0.85) \sigma + (10 \times 1)$$

$$= 10 + 0.35 + 11.9$$

$$\boxed{= 22.25}$$

$$4s z_{eff} = 26 - 22.25$$

$$\boxed{= 3.75}$$

$$\begin{array}{r} 2 \\ 85 \\ \hline 14 \\ 34 \\ \hline 85 \\ \hline 11.90 \end{array}$$

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

$$\alpha = 0.85 \times 2 \\ = 1.7$$

$$Z_{eff} = Z - \alpha \\ = 3 - 1.7 \\ = 1.3$$

4 Be-

$$(1s)^2 \text{ OTTOBLAAR} \\ (2s)^2$$

$$Z_{eff} = 4 - (1 \times 0.35 + 2 \times 0.85) \\ = 4 - 2.05 \\ = 1.95$$

5 B-

$$(1s)^2 (2s, 2p)^3$$

$$Z_{eff} = 5 - (2 \times 0.85 + 2 \times 0.35) \\ = 5 - 2.4$$

$$= 2.6$$

6 C-

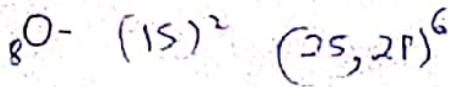
$$(1s)^2 (2s, 2p)^4$$

$$Z_{eff} = 5 - (2 \times 0.85 + 3 \times 0.35) \\ = 5 - 2.75 \\ = 3.25$$

7 N-

$$(1s)^2 (2s, 2p)^5$$

$$Z_{eff} = 7 - (2 \times 0.85 + 4 \times 0.35) \\ = 7 - 3.1 \\ = 3.9$$



$$\begin{aligned} Z_{\text{eff}} &= 8 - (2 \times 0.85 + 5 \times 0.35) \\ &= 8 - (1.7 + 1.75) \\ &= 8 - 3.45 \\ &= 4.55 \end{aligned}$$

$$\begin{aligned} {}^9F^- \quad (1S)^2 \quad (2S, 2P)^7 \\ Z_{\text{eff}} &= 9 - (2 \times 0.85 + 6 \times 0.35) \\ &= 9 - 3.8 \end{aligned}$$

$$\boxed{= 5.2}$$

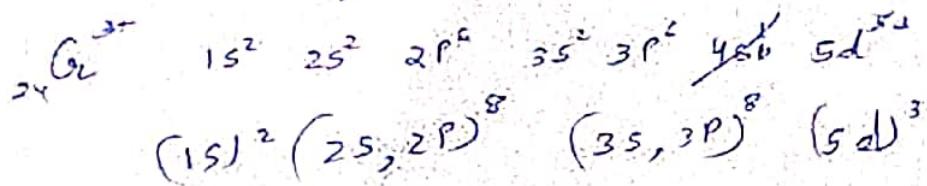
$$\begin{aligned} {}^{10}Ne \quad (1S)^2 \quad (2S, 2P)^8 \\ Z_{\text{eff}} &= 10 - (2 \times 0.85 + 7 \times 0.35) \\ &= 10 - 4.15 \end{aligned}$$

$$\boxed{= 5.85}$$

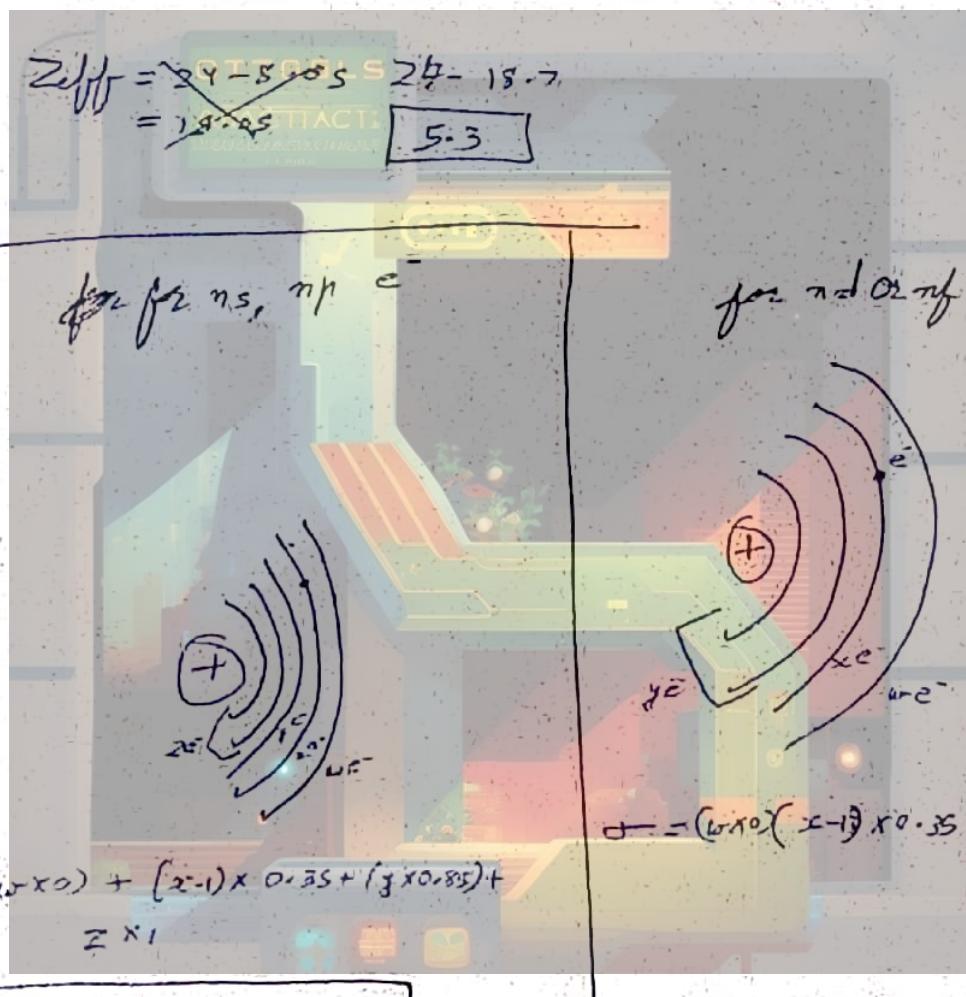
$$\begin{aligned} {}^{11}Na \quad (1S)^2 \quad (2S, 2P)^8 \quad (3S)^1 \\ Z_{\text{eff}} &= 11 - (2 \times 1 + 8 \times 0.85 + 0 \times 0.35) \\ &= 11 - (2 + 6.80 + 0) \end{aligned}$$

$$\boxed{= 2.2}$$

Q Calculate the Z_{eff} of $3d$ e⁻ of Cr³⁺?



$$\begin{aligned}\sigma &= (7 \times 0.35) + (8 \times 0.85) + 2 \\ &= 2.45 + 3.60 + 2 \\ &= 6.05 + 2 \\ &= 18.05\end{aligned}$$



Note:- for $1s^2$ e⁻ configuration :-

$$\sigma = 0.3$$

$$\text{eg. He} - 1s^2$$

$$o = 0.3$$

$$Z_{eff} = 2 - 0.3$$

$$= 1.7$$

$$\text{Li}^+ - 1s^2$$

$$o = 0.3$$

$$Z_{eff} = 3 - 0.3$$

$$= 2.7$$

$$\text{Ni} - 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^8$$

$$\sigma = 80.3$$

$$Z_{eff} = 28 - 0.3$$

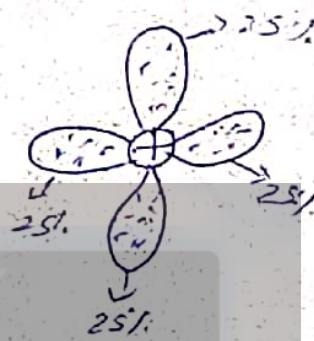
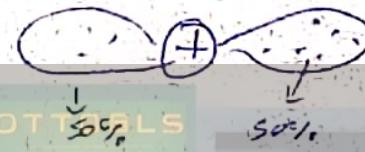
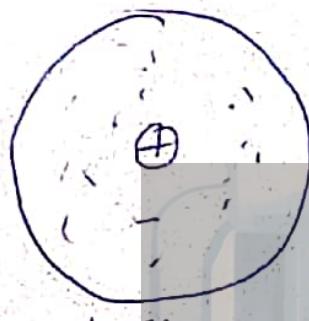
$$= 27.7$$

Note:- For Helium, $Z_{eff} = 1$

Penetration Power - It is the ability of an orbital to attract e^-

order - $nS > nP > nD > nF$

(how easy can electron approach to nucleus)



Shielding Power / Screening power :-

order - $nS > nP > nD > nF$

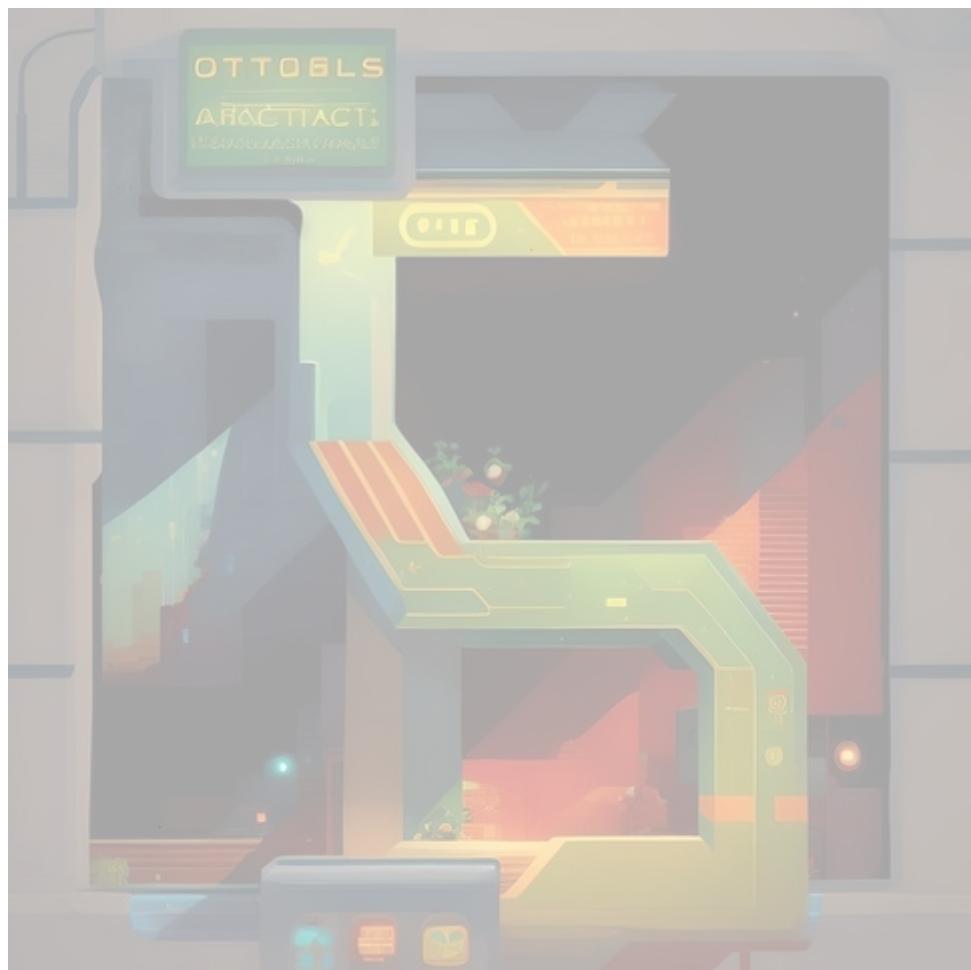




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(55)





Periodic Table & Periodic Properties

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

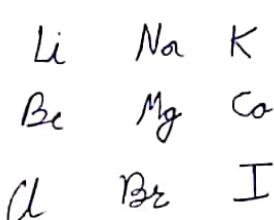
Development of periodic table:-

1. Lavoisier - classification of elements
2. Döbereiner's - Triad rule
3. Newlands - Rule of octave
4. Lothar - Meyer curve
5. Mendeleev's - Periodic table
6. Moseley's - Experiment off and modern periodic table.

1. Lavoisier Classification - firstly he classified elements into 2 categories



2. Döbereiner's Triad rule - He made the group of 3 elements having similar chemical properties called ~~triple~~ triad.
→ In Döbereiner's triad, atomic mass of central element is nearly equal to average atomic mass of 1st & 3rd element



Li	Na	Mg
7	23	39
$\frac{7+39}{2} = \frac{46}{2} = 23 = \text{Na}$		

K Rb Cs

S Se Te

Cl	Br	I
35.5	81.25	127
	81.25	

$$\begin{array}{c} \text{Ca Sr Ra} \\ \boxed{\text{F Cl Br}} \rightarrow \text{not following} \end{array}$$

$$\frac{127 + 35.5}{2} = \frac{162.5}{2} = 81.75 \approx \text{Br}$$

F	Cl	Br
19	35.5	81.25 80

$$\frac{80 + 19}{2} = 49.5 \neq \text{Cl}$$

Q1. Multi Correct

If three element $X^{20} Y^{42} Z^{20}$
 follow Dobereiner's triad. find atomic mass of Z

- A) 64 B) 38 C) 31 D) 29

3. Newland's law of octave - He arranged elements in order of increasing atomic mass and found that every eighth element showing similar properties to the 1st element like musical notes.

1	2	3	4	5	6	7	8
Sa	Be	Ca	Ma	Pa	Dha	Ni	
Sa							

Li	Be	B	C	N	O	F
Na	Mg	Al	Si	P	S	cl
K	Ca	(Xc)				

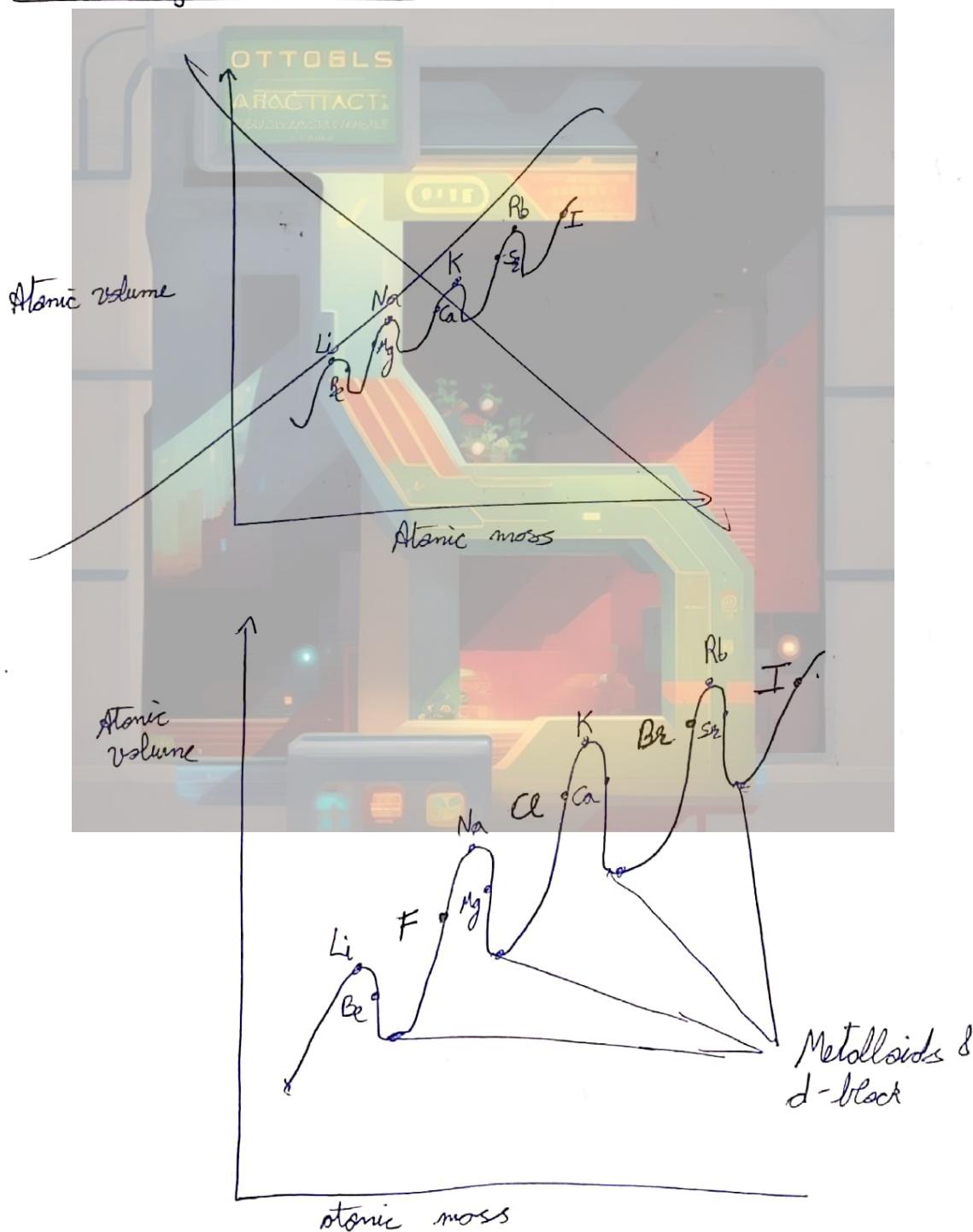
* * * Noble gases were not discovered yet

merit:- He classified the known elements in tabular form.

demerit:- This rule is valid only upto calcium because after calcium due to filling of (d) electrons their orbitals is a difference of 18^{th} element instead of 8^{th} element.

- discovery of noble gases failed the ~~rule~~ rule.

4. Lanthanide Moyer's curve -



Observation of curve - The following points can be observed

- similar properties of elements occupied similar portion of the curve.
- Alkali metals (Li, Na, K, Rb, Cs) occupied peak portions of the curve.
- Alkaline earth metals (Be, Mg, Ca, Sr) occupied the descending portion of the curve.
- Halogens (F, Cl, Br, I) occupied ascending portion of the curve.
- Metals & transition elements occupied lower portions of the curve

Drawbacks → It is highly complex when further ~~metals~~ elements are plotted on the curve.

5. Mendeleev's Periodic Table :- Mendeleev's Periodic Table is based on Mendeleev's periodic law, which states that physical and chemical properties of elements are periodic function of their atomic mass.

1	Zero Group
2	He
3	Ne
4	Ar
5	Kr
6	Xe
7	Rn

1. Around 63 elements were arranged in order of increasing atomic weight into 7 horizontal rows (called periods)

and 8 vertical columns (called groups).

2. Each group except 8th group was further divided into 2 sub-

groups

3. Subgroup IA contains main elements while subgroup B contains transition elements.

4. At that time noble gases were not discovered and when they were discovered an extra column called zero group was added.

5. The elements of some group have similar properties.

~~Merits~~ - 1. Old known all known elements were arranged.
Merits 2. He left spaces for undiscovered elements and the properties of these elements were accurately predicted.

Mendeleev's Name

Eka-Aluminium (Al)

Eka-Silicon (Si)

Eka-Boron (B)

Eka-Manganese (μ n)

Modern Name

Calcium (Ca)

Germanium (Ge)

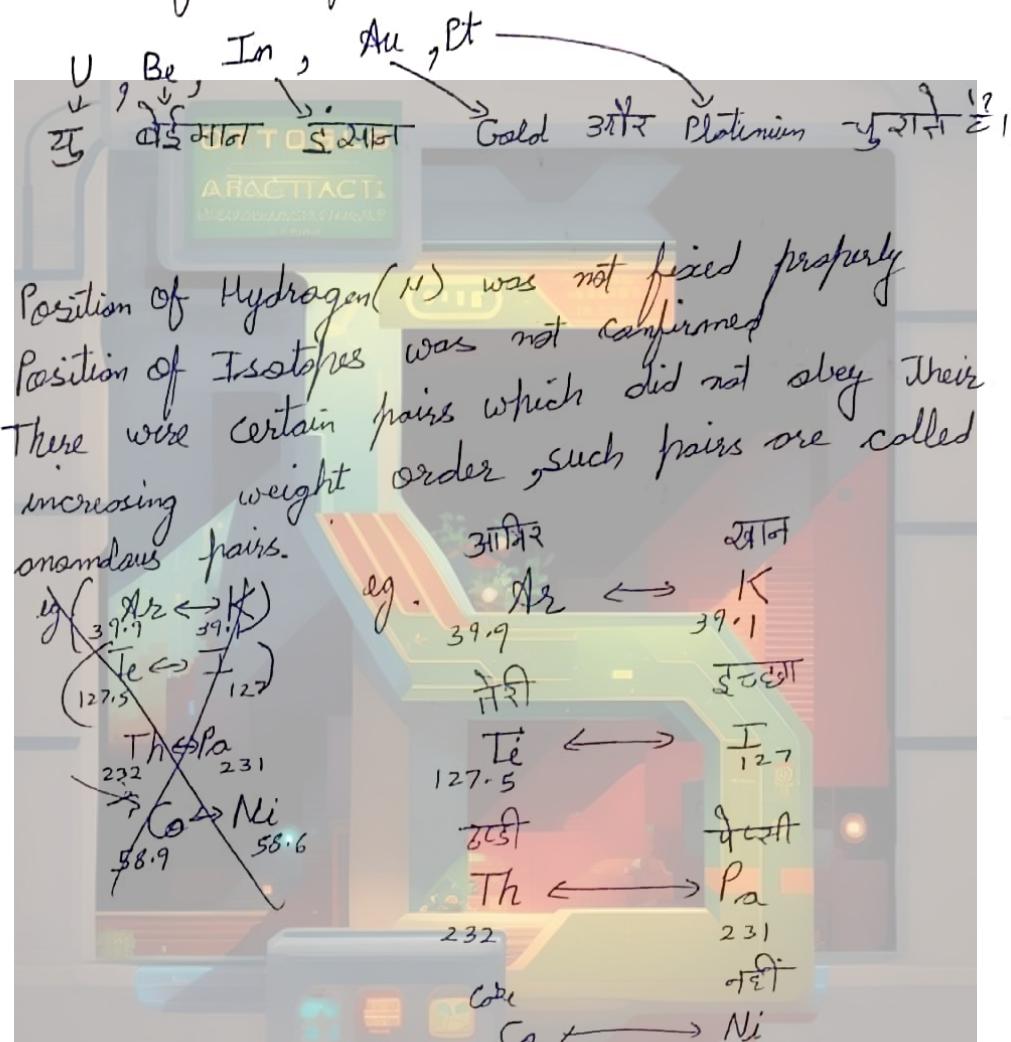
Scandium (Sc)

Technetium (β Tc)

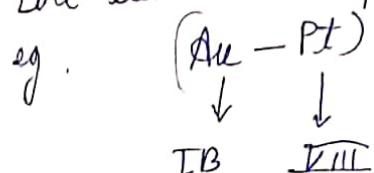
Si - Ge 2-
 शिला - ग्लास
 He - Ge
 31 पर 31

B - Sc Mn-Tc
 B. Sc नहीं M. Tech वै

3. Correction of doubtful atomic weights



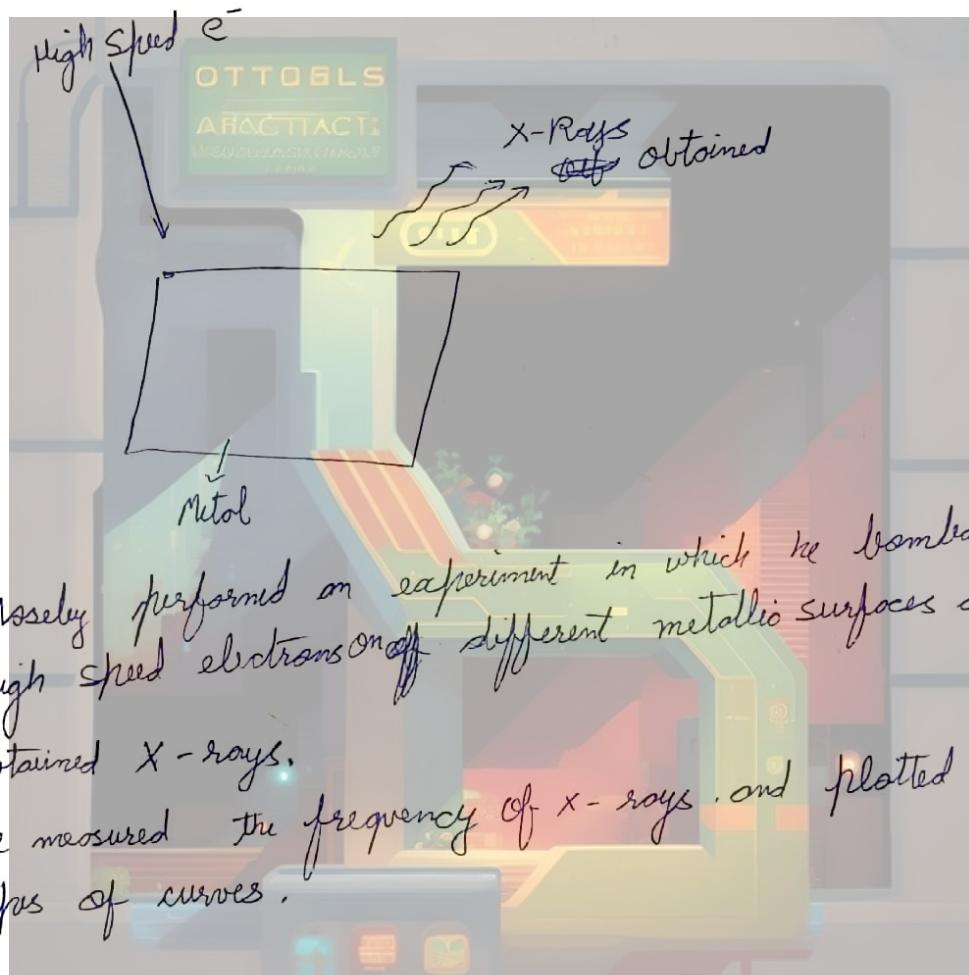
- Like elements were placed in different groups



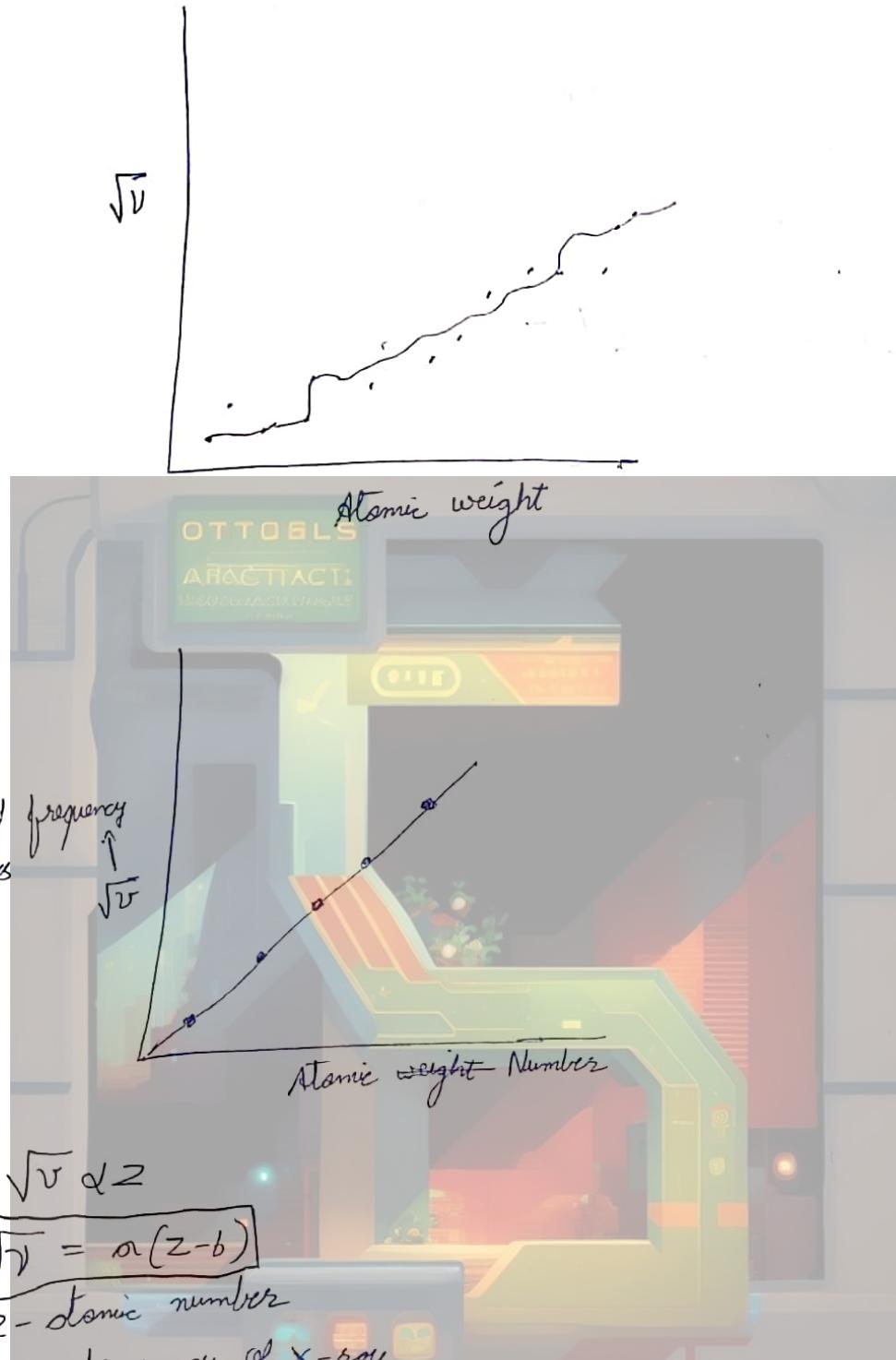
- Unlike elements were placed in same group.

g.	I-A	
Li		Highly reactive
Na		
K	IB	
Rb		Cu
Cs		Ag } coinage metals
		Ag } less reactive
		Au

Moseley's Experiment -



- Moseley performed an experiment in which he bombarded high speed electrons on different metallic surfaces and obtained X-rays.
- He measured the frequency of X-rays and plotted various types of curves.



→ Moseley's experiment showed that the fundamental property of an element is not atomic mass but atomic number.

Q Identify the element for which Z-ray frequency is
625 Hz. given ($a=1$, $b=2$)

$$\sqrt{625} = 25$$

$$25 = \alpha(z-1)$$

$$25 = 1(z-1)$$

$$25 = z-1$$

$$\cancel{z=26}$$

$$z=27$$

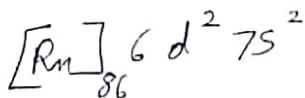
Cobalt (Co)

Modern Periodic Table

- It is based on modern periodic law which states that the physical and chemical properties of elements are the periodic function of their atomic number.
- It consists of total 18 vertical columns called groups, and 7 horizontal rows called periods.

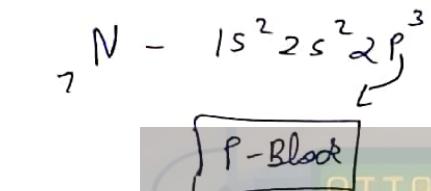
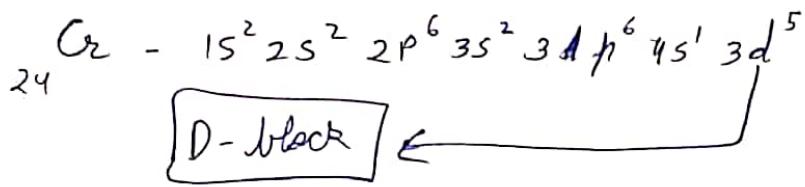
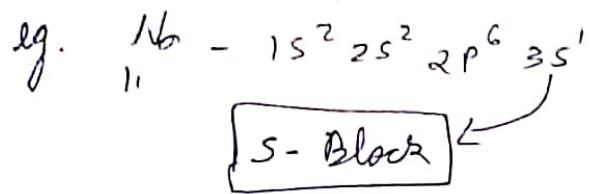
Deciding the block of an element -

- There are 4 blocks in a periodic table which are S, P, D & F.
- To decide the block we first write the electronic configuration of the element.
- The type of subshell in which last electron enters will give the information about block of that element. (Exception - Th (Thorium))



last electron enters in d subshell but Th is F block element

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Group I (IA)
Config.: ns^1 (valence shell $e^- = 1$)

$1s^1$	H	Non-metal	सू
$2s^1$	Li		ली
$3s^1$	Na		ना
$4s^1$	K		की
$5s^1$	Rb		रब
$6s^1$	Cs		सी
$7s^1$	F ₂		फॉर्सियाद

Alkali metals
(Highly reactive)

Group II (IIA)
Config: ns^2 (valence shell $e^- = 2$)

$2s^2$	Be	देटा
$3s^2$	Mg	मॉग्ने
$4s^2$	Ca	कॉर
$5s^2$	Sr	सैर्कुलर
$6s^2$	Ba	बॉय
$7s^2$	Rs	रासी

Alkaline earth metals

Note - Group 1 of Group 15 are s-block elements

Group 13 (III-A) (valence shell $e^- = 3$)

config - $nP^1 nS^2 nP^1$

$2P^1 B$
 $3P^1 Al$
 $4P^1 Ga$
 $5P^1 In$
 $6P^1 Tl$
 $7P^1$

\rightarrow Boron family

বৰণ পৰিয়াল

বৰণ

আল

গান্ধি

In

টেলা

Group 14 (IV-A) (valence shell $e^- = 4$)

config - $nS^2 nP^2$

$2P^2 C$
 $3P^2 Si$
 $4P^2 Ge$
 $5P^2 Sn$
 $6P^2 Pb$

- Carbon Family

কার্বন
সিলিস
জিনি
সুনো
পৰম্পৰা

কার্বন

সিলিস

জিনি

সুনো

পৰম্পৰা

Group 15 (V-A) (valence shell $e^- = 5$)

config - $nS^2 nP^3$

$2P^3 N$
 $3P^3 P$
 $4P^3 As$
 $5P^3 Sb$
 $6P^3 Bi$

Nitrogen family (pnictogens)

Nepal

Pakistan

Australia

Sov

Bikhari

Group 16 (VI-A) (valence shell $e^- = 6$)

config - $nS^2 nP^4$

$2P^4 O$
 $3P^4 S$
 $4P^4 Se$
 $5P^4 Te$
 $6P^4 Po$

Oxygen family (chalcogens)

old

style

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(69)

Group - 17 (VIIIA) (valence shell $e^- = 7$)

config - $ns^2 np^5$

$2P^5$	F	Halogen family	$\frac{F}{F}$
$3P^5$	Cl		$\frac{Cl}{Cl}$
$4P^5$	Br		$\frac{Br}{Br}$
$5P^5$	I		$\frac{I}{I}$
$6P^5$	At		$\frac{At}{At}$

Sulphur / Sulfur

Group - 18 (zero group) (valence shell $e^- = 8$)

config - $ns^2 np^6$ (valence shell $e^- = 8$)

~~Group~~ 1s² He (s-block element)

$2P^6$ Ne
 $3P^6$ Ar
 $4P^6$ Kr
 $5P^6$ Xe
 $6P^6$ Rn

Noble gases
(inert gases) $\frac{He}{He}$

$\frac{Ne}{Ne}$
 $\frac{Ar}{Ar}$
 $\frac{Kr}{Kr}$
 $\frac{Xe}{Xe}$
 $\frac{Rn}{Rn}$

Note:-

1. Group 13, 14, 15, 16, 17 & 18 are known as p-block elements.

1. Group 13, 14, 15, 16, 17 & 18 are known as p-block elements except (Helium)

2. Helium is s-block element.

3. Helium is known as noble gases/inert gases.

4. Except group 18 elements all s-block and p-block

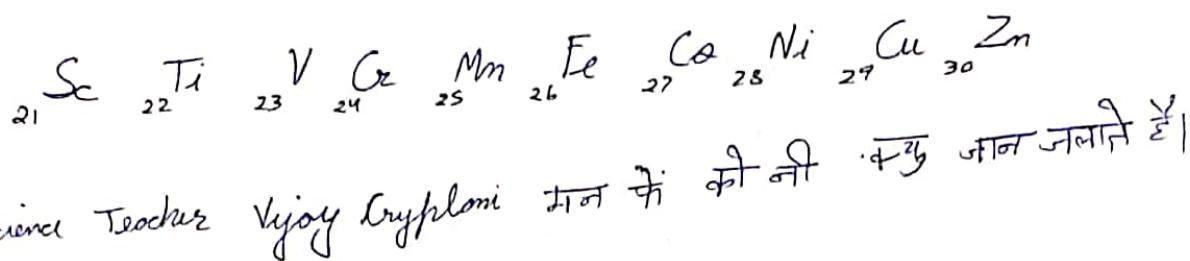
elements are known as normal or representative elements.

Representative elements - Outermost shell of these elements is incomplete.

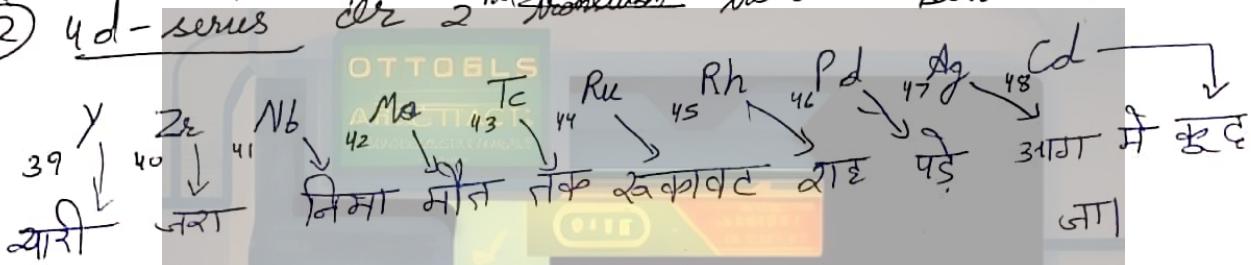
The number of e^- in outermost shell is less than 8.

D-block elements :-

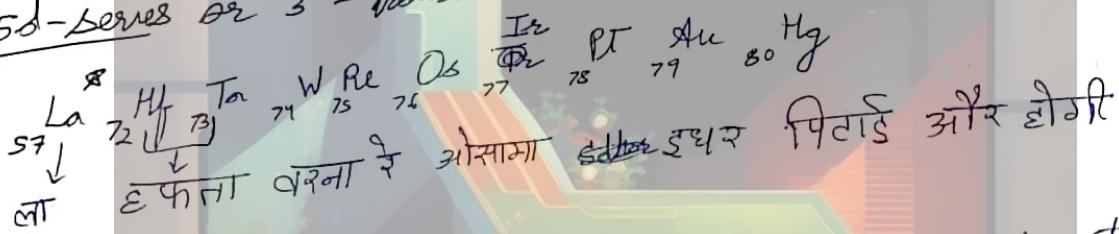
① 3d-series or Ist transition series elements.



② 4d-series or 2nd transition transition series elements.



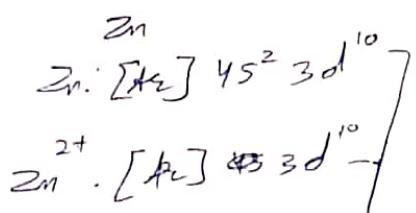
③ 5d-series or 3rd transition elements.



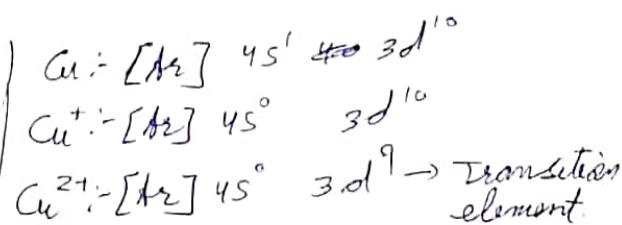
Note:-
1. Elements of these blocks are called transition elements

2. Transition elements are those elements which have incompletely filled d subshell either in their ground state or any of their cationic state found in nature

3. Zinc, Cadmium (Cd) & Mercury (Hg) are not transition elements yet they are d block elements.



not Transition element



F-Block elements

1. 4f-block element /, at inner transition elements / Lanthanides elements /

Lanthanoids / Rare earth metals

D-block element

${}_{57}^{7}\text{La}$	${}_{58}^{7}\text{Ce}$	${}_{59}^{7}\text{Pr}$	${}_{60}^{7}\text{Nd}$	${}_{61}^{7}\text{Pm}$	${}_{62}^{7}\text{Sm}$	${}_{63}^{7}\text{Eu}$	${}_{64}^{7}\text{Gd}$
लान्थनियम	सीरियम	प्रस्ट्रियम	नेडियम	प्रमेरियम	स्मिरियम	यूरियम	गोडियम

Diagram showing the lanthanide series from ${}_{90}^{7}\text{Th}$ to ${}_{71}^{7}\text{Lu}$. The elements are labeled with their atomic numbers and names in English and Hindi. The diagram also includes the labels 'OTTOSLS' and 'ACTINIDES'.

2. 5f-block elements / 2nd inner transition elements / & Actinides Elements / Actinoids.

Diagram showing the actinide series from ${}_{89}^{89}\text{Ac}$ to ${}_{99}^{99}\text{Es}$. The elements are labeled with their atomic numbers and names in English and Hindi. The diagram also includes the label 'D-block element'.

Note:-

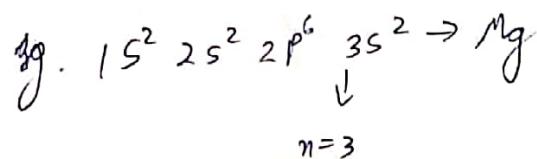
1. Lanthanum & Actinium (La & Ac) are D-block elements

2. Lanthanide series elements & Actinide series elements are F-block elements.

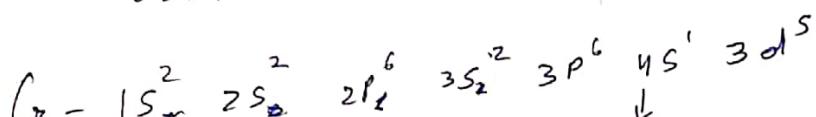
3. A Lanthanide series starts from ${}_{57}^{7}\text{Ce}$ to ${}_{71}^{7}\text{Lu}$

4. Actinide series starts from ${}_{90}^{90}\text{Th}$ to ${}_{103}^{103}\text{Lr}$

Description of Period - The highest value of n in electronic configuration of an element will give its period.



3rd Period



4th period

No. of elements in a specific period.

Period	Subshell	no. of element	Name
1 st Period	1s	2	Shortest
2 nd Period	2s 2p	10	Short
3 rd Period	3s 3p	10	Short
4 th Period	4s 3d 4p	18	long
	5s 4p	18	long
5 th Period	6s 4f 5d 6p	32	longest
6 th Period	7s 5f 6d 7p	32	elongated
7 th Period			

Note:- 1. Any Period start from 'ns' & end with 'np'

ms... $(n-3)g$ $(n-2)f$ $(n-1)d$ n^p till n^l

Q1. If each orbital is occupied by max $3e^-$, find total no. of elements in 2nd period.

In 2nd period $2s, 2p$

$$2s - \boxed{111} - 3 e^- \rightarrow 3 \text{ elements}$$

$$2p - \boxed{111111} - 9 e^- \rightarrow 9 \text{ elements}$$

Total elements $\rightarrow \boxed{12 \text{ elements}}$ find total no. of elements in 2nd period

Q2. If each orbital occupied by max $3e^-$, find total no. of elements in 6th period

6th Period - $6s 4f 5d 6p$

$$6s \rightarrow 3 e^-$$

$$4f \rightarrow 14 + 7 \rightarrow 21 e^-$$

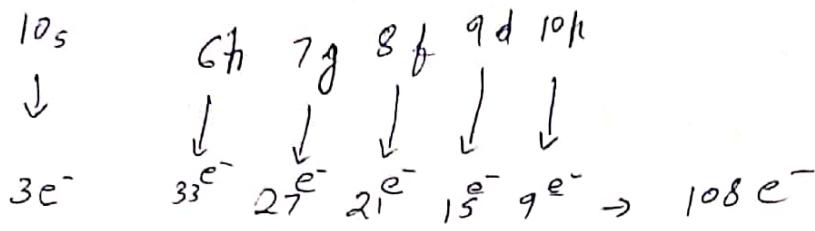
$$5d \rightarrow 10 + 5 \rightarrow 15 e^-$$

$$6p \rightarrow 6 + 3 \rightarrow 9 e^-$$

$$\text{Total } e^- = 48$$

$$\boxed{\text{Total elements} = 48}$$

Q3. If each orbital occupied by max $3 e^-$, find no. of elements in 10th period.



$\boxed{\rightarrow 108 \text{ elements}}$

Q4. find atomic no. of element present in 8th period group 17.

$$118 + 17 = 135$$

7th period total = 118 elements

~~17th group = 17~~

~~$118 + 17 = \boxed{+ 35^{\text{th}} \text{ element}} + 18$~~

~~= 153 element~~

18th group \rightarrow 50th element of period
17th group \rightarrow 49th element of period

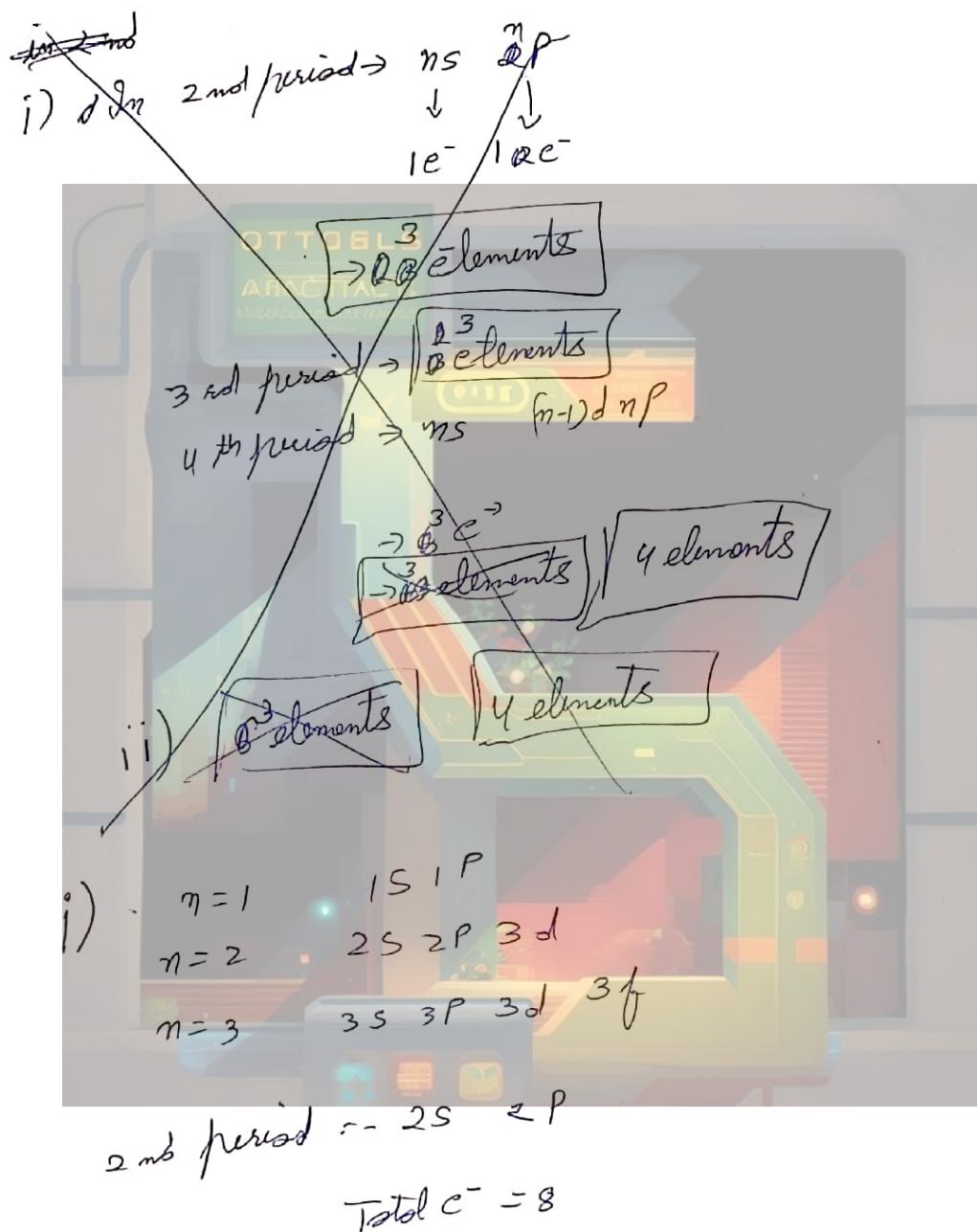
$$118 + 49 = 167$$

$\boxed{Z = 167}$

Q4. In other universe if all quantes nos. are same except l , $l = 0 \text{ to } n$

i) number of elements in period 2, 3, 4

ii) elements in 3rd shell



3rd period \rightarrow 3s 2d 3p

\rightarrow 18 elements

4th period \rightarrow 3s 3p
4s 3d 4p

\rightarrow 18 elements

ii)

$n=3$

\rightarrow 3s 3p 3d 3f
 $\downarrow \quad \downarrow \quad \downarrow \quad \downarrow$
2 6 10 14

\rightarrow 32 elements

Q 5. In other universe, if O. No. same except

$l \rightarrow Oto(n+1)$

$m \rightarrow -(l+1) \text{ to } +(l+1)$

$m_s \rightarrow -\frac{1}{2}, 0, +\frac{1}{2}$

Then, calculate

no. of elements in 2nd period

elements in 2nd shell

no. of

2nd period \rightarrow 2s 1d 2p
 $\downarrow \quad \downarrow \quad \downarrow$
2s orbital 7 5
 $\downarrow \quad \downarrow$
 $3 \times 3 = 9 e^-$ $3 \times 7 = 21 e^-$ $3 \times 5 = 15 e^- = 45 e^-$

\rightarrow 45 elements

2nd Shell $\rightarrow n=2$

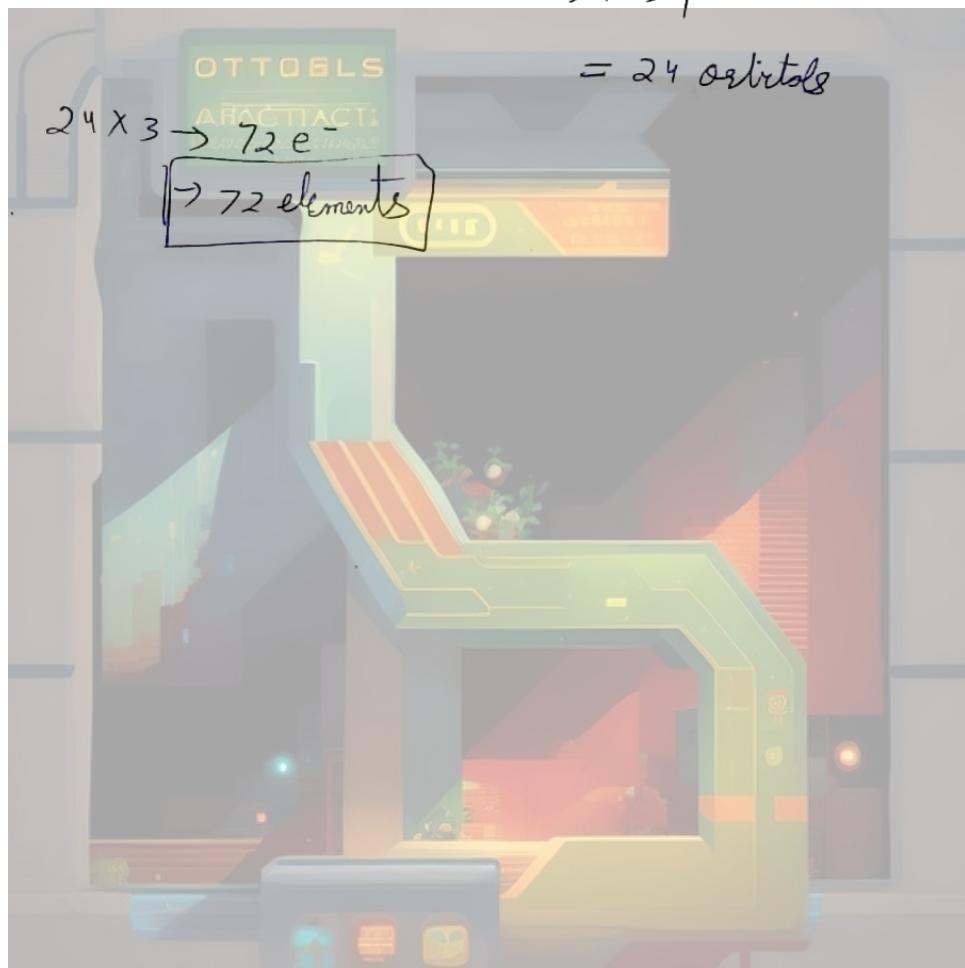
$l = 0(s)$
 $1(p)$
 $2(d)$
 $3(f)$

$m = -1, 0, 1 \rightarrow 3$

$-2, -1, 0, 1, 2 \rightarrow 5$

$-3, -2, -1, 0, 1, 2, 3 \rightarrow 7$

$-4, -3, -2, -1, 0, 1, 2, 3, 4 \rightarrow 9$



Some Commonly used terms:-

1. Post Lanthanides elements - After lanthanide elements including (after Lu, $Z=71$) or Hafnium ($Z=72$)

not included onwards elements are known as post lanthanide elements.

2. Trans Uranic elements - After Uranium element, atomic no. 92

excluded ($Z=92$) or (Np , $Z=93$) are known as Trans \uparrow Uranic elements.

3. Typical elements - 2nd & 3rd period elements are known as typical elements

Classification of elements according to Block :-

1. S-Block elements - last e- enters in ~~s-block~~

S-Subshell,

General electronic Config nS^{10-2}

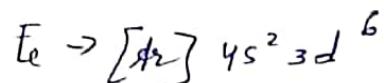
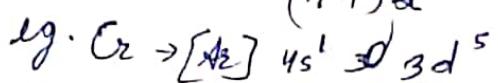
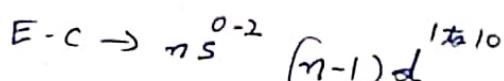
nS^1 or nS^2

e.g. Li - $1S^2 2S^1$

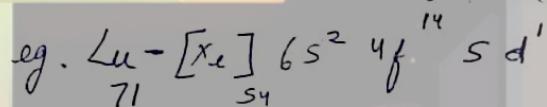
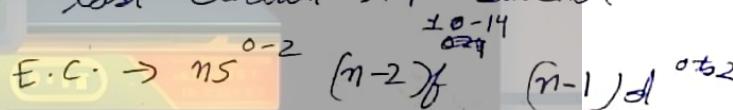
Mg : $[Ne] 3S^2$

2. P-Block elements - lost electron in P-Subshell
 $E.C \rightarrow ns^2 \xrightarrow{\text{lose } e^-} n_p^1$
 eg. $N \rightarrow [He] 2s^2 2p^3$

3. D-Block elements - lost electron in D-Subshell



4. F-block elements - lost electron in F-Subshell

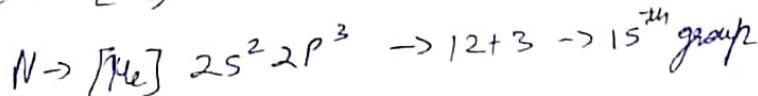
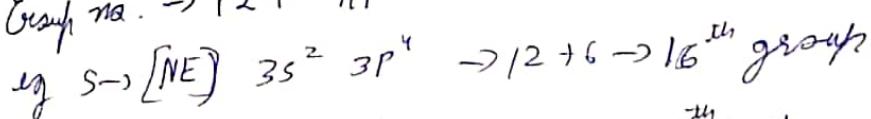


Trick to estimate group number of an electron

- To decide the group of an element, first decide the block of the element.
 - S-Block -
 (except Helium) $ns^1 \rightarrow$ 1st group H, Li, N, K, Rb, Cs, Fr
 $ns^2 \rightarrow$ 2nd group Be, Mg, Ca, Sr, Ba, Ra
- no. of electron in valence shell = Group no. (for S-Block)

→ P-Block -

$\text{Group no.} \rightarrow 12 + np e^-$



③ Group no. of D-block elements

Group no. \rightarrow no. of nsc^- + no. of ~~s⁻¹~~ $(n-1)d\ e^-$

e.g. Cr - [Ar] $4s^1 3d^5$

group $\rightarrow 1+5 \rightarrow 6^{\text{th}}$ group.
 \rightarrow D-block.
 \rightarrow 4th period.

e.g. 2. Fe \rightarrow [Ar] $4s^2 3d^6$

group $\rightarrow 6+2 \rightarrow 8^{\text{th}}$ group.
 OTTO \rightarrow D-block.
 ARCTIC \rightarrow 4th period.

④ Group no. of F-block - Group - 3 always

Q1. Atomic no. 32, 56, 27, 63, 2, 83

32 \rightarrow [Ar] $4s^2 3d^{10} 4p^2$ $\left| 27 \rightarrow [Ar] 4s^2 3d^7 \right.$
 \rightarrow P-block
 \rightarrow 4th period
 \rightarrow 14th group

56 \rightarrow [Xe] $6s^2$
 \rightarrow S-block
 \rightarrow 6th period
 \rightarrow 2nd group

63 \rightarrow [Xe] $6s^2 5d^1 4f^7$
 \rightarrow F-block
 \rightarrow 6th period
 \rightarrow 30th group

2 \rightarrow $1s^2$
 \rightarrow S-block
 \rightarrow 1st period
 \rightarrow 1st group

83 \rightarrow [Xe] $6s^2 4f^{14} 5d^{10} 6p^3$
 \rightarrow P-block
 \rightarrow 6th period
 \rightarrow 15th group.

IUPAC Nomenclature of elements

(atomic num. $> 10^6$)

→ All names must end with "ium"

Bi + ium → Biium

Pent + ium → Pentium

Digit	name	Abbreviation
0	Nil	n
1	Un	u
2	Bi	b
3	Tri	t
4	Quad	q
5	Pent	p
6	Hex	h
7	Sept	s
8	Oct	o
9	En	e
+0		

Ex ① - $Z=118$ (Uuo)
with Ununoctium

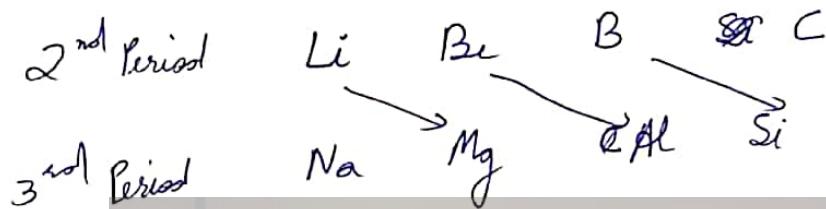
④ $Z=115$ (Uup)
Ununpentium

② $-Z=104$ (Uumg)
Unnilquadium

③ $Z=110$ (Uun)
Ununnilium

Diagonal Relationship

→ Properties of elements of 2nd Period resembles with the element of 3rd Period. These resembles between properties of 2nd and 3rd period is called diagonal relationship.



Li & Mg show similar properties -

eg. $\text{Li}_2\text{CO}_3 \xrightarrow{\Delta} \text{Li}_2\text{O} + \text{CO}_2$

$\text{MgCO}_3 \xrightarrow{\Delta} \text{MgO} + \text{CO}_2$

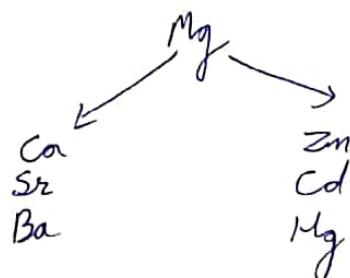
$\text{Na}_2\text{CO}_3 \xrightarrow{\Delta} \text{No reaction}$

BeO > Both are Amphoteric
 Al_2O_3

Bridge Element

Mg of 3rd period is bridge element

Be



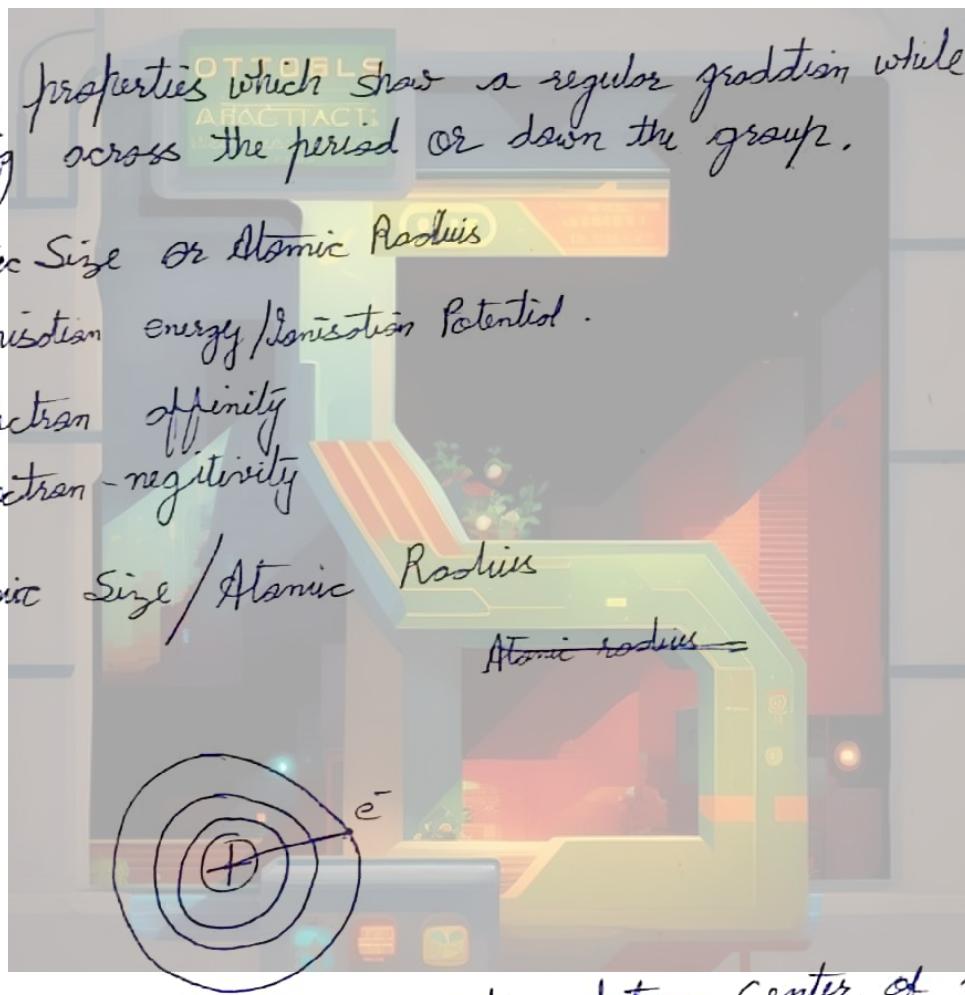
→ The properties of Bridge elements are somewhat mixed of the elements of 2 sub groups as ~~as~~
magnesium shows similarities with alkaline earth metals on one hand and with zinc metals on other hand.

Periodic Properties

→ Those properties which show a regular gradation while moving across the period or down the group.

- 1. Atomic Size or Atomic Radius
- 2. Ionisation energy/Ionisation Potential
- 3. Electron affinity
- 4. Electron-negativity

→ Atomic Size / Atomic Radius



Atomic Size - It is the distance between center of nucleus and the outermost shell electron.

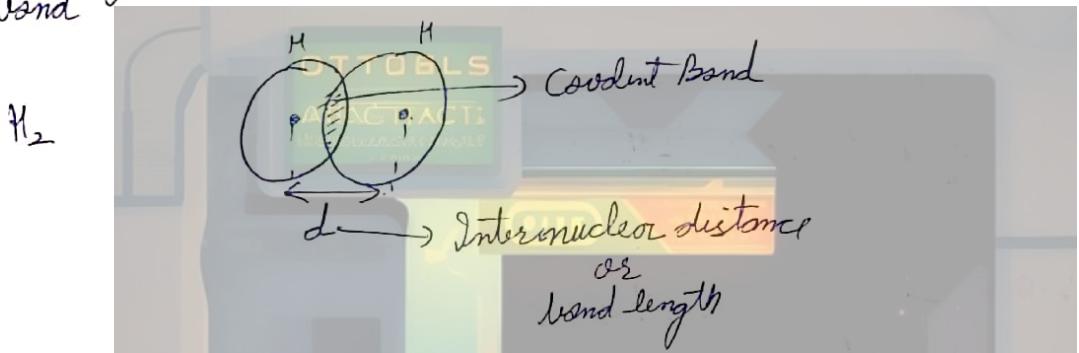
→ It is impossible to calculate exact value of radius or actual value of atomic size

→ Since we are calculating the radius of atoms in different bonded state (covalent bond → covalent radius → for non-metal)
 Metallic bond → Metallic radius → (for metal)
 Vanderwaal's bond → Vanderwaal radius
 Ionic bond → Ionic radius

↓
 (generally held gases)

Covalent Radius →

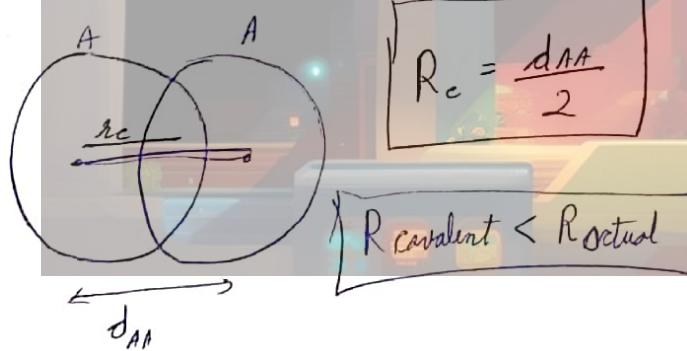
→ It is defined for those elements which are bonded with covalent bond



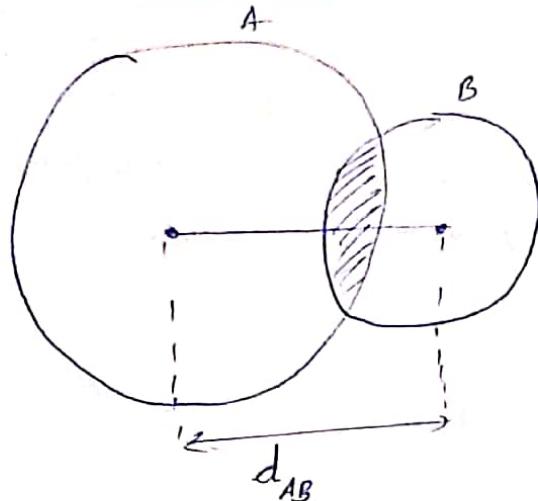
→ Bond length - It is interatomic distance between two atoms in a bonded state.

Covalent radius

① Homo-diatomic radius (H_2 , Cl_2 , O_2 etc.)



② Hetero-diatomic molecule (H-F, H-Cl)



Schoenheimer - Stevensons Equation -

$$d_{AB} = R_A + R_B - 0.09 |EN_A - EN_B|$$

R → Radius

$$d_{AB} = R_A + R_B - 0.09 |EN_A - EN_B|$$

R_A → radius of A

R_B → radius of B

EN_A & EN_B → Electronegativity of A & B on Pauling scale

d_{AB} → Bond length between A & B

Distance in \AA

$$1 \text{ \AA} (\text{Angstrom}) = 10^{-10} \text{ m}$$

$$1 \text{ pm} (\text{picometer}) = 10^{-12} \text{ m}$$

$$1 \text{ nm} (\text{nano meter}) = 10^{-9} \text{ m}$$

$$1 \mu\text{m} (\text{micro meter}) = 10^{-6} \text{ m}$$

$$1 \text{ mm} (\text{mille meter}) = 10^{-3} \text{ m}$$

$$d_{AB} = R_A + R_B - 9 |EN_A - EN_B|$$

(In pm)

Q Calculate bond length of HCl molecule, if bond length of H_2 is 2 \AA° & Cl_2 molecule is 10 \AA° ; given EN_H is 2.1 & EN_{Cl} is 3.

$$d_{HH} = 2\text{ \AA}^\circ$$

$$d_{ClCl} = 10\text{ \AA}^\circ$$

$$d_{HH} = R_H \times 2$$

$$\frac{10}{2} = R_{Cl}$$

$$2\text{ \AA}^\circ = R_H \times 2$$

$$R_{Cl} = 5\text{ \AA}^\circ$$

$$R_H = 1\text{ \AA}^\circ$$

$$d_{HCl} = 1 + 16 - 0.09(2.1 - 3)$$

$$= 7 - 0.09(0.9)$$

$$= 7 - 0.081$$

$$= 6.919$$

$$- 0.081$$

$$56.919\text{ \AA}^\circ$$

Q2. Calculate bond length of AB if $R_A = 15\text{ pm}$
 $R_B = 20\text{ pm}$

$$EN_A = 2$$

$$EN_B = 1.5$$

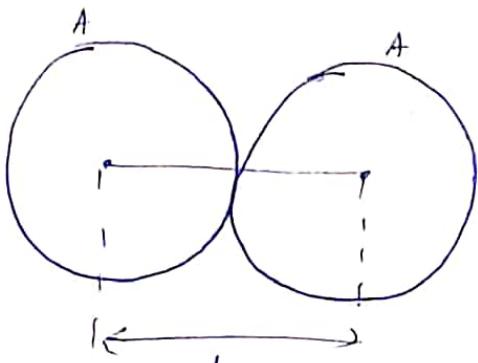
$$d_{AB} = 15 + 20 - 9(2 - 1.5)$$

$$= 35 - 9(0.5)$$

$$= 35 - 4.5$$

$$= 30.5\text{ fm}$$

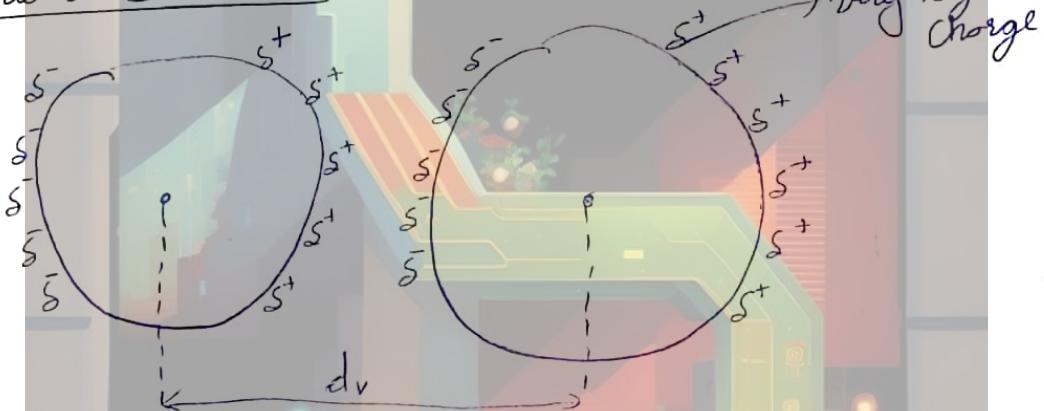
② Metallic Radius/Crystal Radius - It is half the internuclear distance between two identical atoms bonded by a metallic bond.



$$R_{\text{metallic}} = \frac{d}{2}$$

$R_{\text{metallic}} \approx R_{\text{actual}}$

③ Vanderwaal's Radius - due to weak forces



$$R_v = \frac{d_v}{2}$$

$$R_v > R_{\text{actual}}$$

$$R_v > R_m > R_c$$

↓ vanderwaal ↓ metallic ↓ covalent

Note - Generally for inert gases or noble gases, Vanderwall radius is calculated.

factors affecting atomic size -

①

$$\text{Size} \propto \frac{1}{Z_{\text{eff}}}$$

$Z_{\text{eff}} \uparrow$ Size \downarrow

②

$$\text{Shell no.} \propto \text{Size}$$

Left Top across the period
 $Z_{\text{eff}} \uparrow$, size \downarrow

Bottom $n \uparrow$
size \uparrow (in general)

Right

for S or P block -

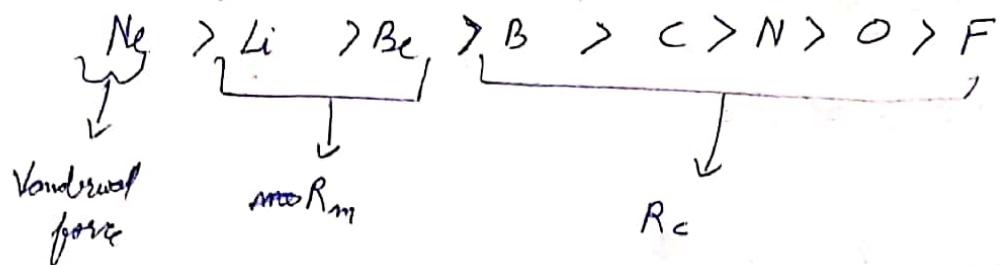
(In Period)

1st period - H > He (expected)

He > H (actual)

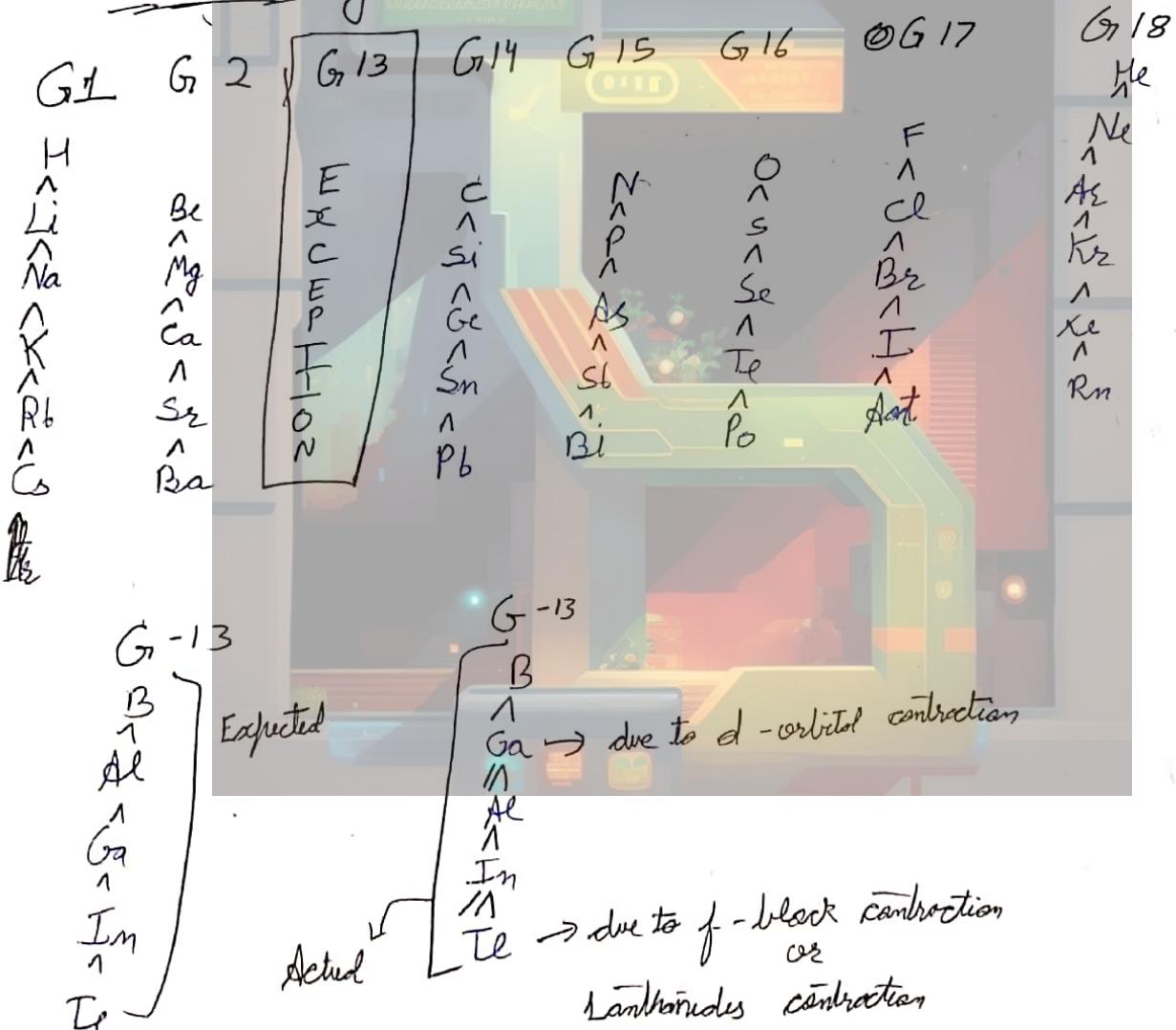
\downarrow
Vanderwall
radius

2nd Period - $\text{Ne} > \text{Li} > \text{Be} > \text{B} > \text{C} > \text{N} > \text{O} > \text{F} > \text{Ne}$ *Nodal process*



3rd Period - $\text{Ar} > \text{Na} > \text{Mg} > \text{Al} > \text{Si} > \text{P} > \text{S} > \text{Cl}$

Variation along the group -



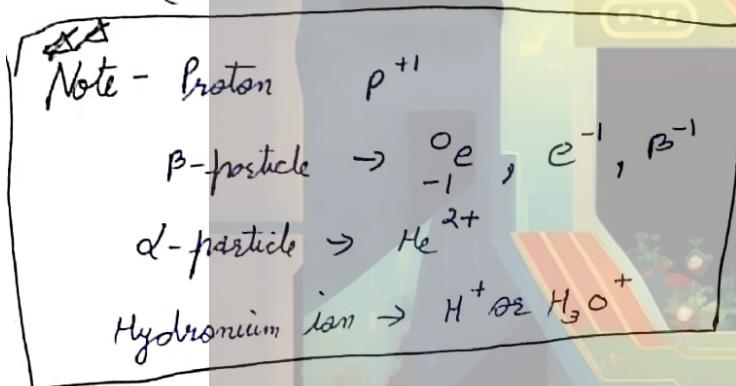
D-Contraction - Due to poor shielding of D-subshell or D-electrons, effective nuclear charge increases and hence size of atom decreases.

Lanthanoid Contraction - Due to poor shielding of F-electrons, effective nuclear charge increases and thus, size of atom decreases.

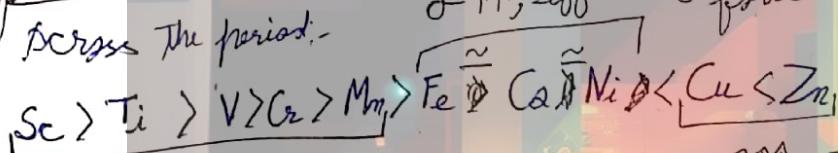
H.W. 21/04/2025

Topic 1-4

$$\left. \begin{array}{l} \text{Size of atom} = 10^{-10} \text{ m} \\ \text{nucleus} = 10^{-14} \text{ m} \\ p/\pi/e^{-1} = 10^{-15} \text{ m} \end{array} \right\} \text{approx}$$



for D-block -
across the period - $\sigma \uparrow, Z_{eff} \uparrow$ (both attraction force dominant, force of repulsion)

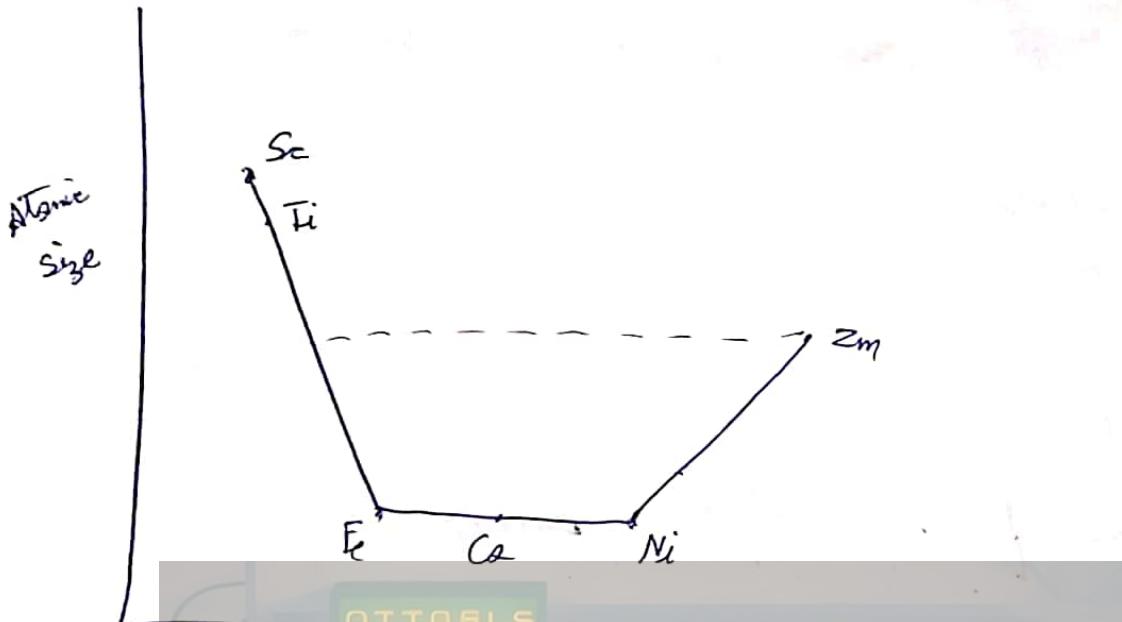


$\sigma \uparrow, Z_{eff} \uparrow$

Size \downarrow

(attraction force dominant)

$\sigma \uparrow, Z_{eff} \uparrow$ (repulsion force dominant)



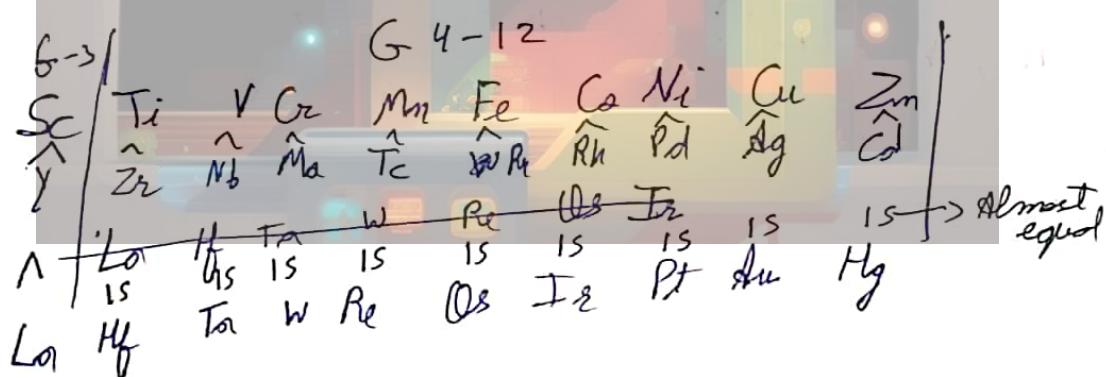
(~~more~~)

variation along the groups

Group - 3 - $3d < 4d < 5d$

$Sc < Y < La$

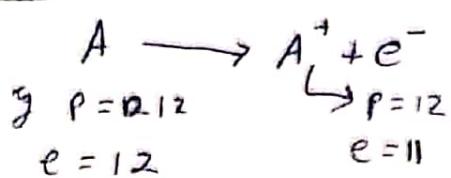
Group - 4 to 12 - ($3d < 4d \approx 5d$)



2 3

Ionic Radius (Cations or Anions)

Cation -



Attraction \uparrow , Size \downarrow

 Size of atom > Size of cation

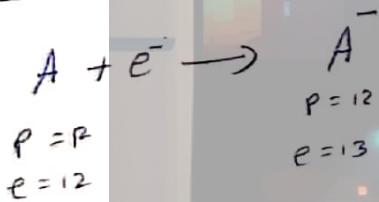
Radius order:-

$$[A > A^+ > A^{2+} > A^{3+}]$$

Eg

$$\begin{aligned} Na &> Na^+ \\ Mg &> Mg^{2+} \\ Al &> Al^{3+} \end{aligned}$$

Anion -



Attraction \downarrow , Size \uparrow

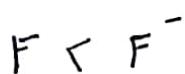
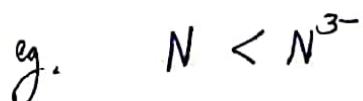
 Size of Atom < Size of Anion

Radius order

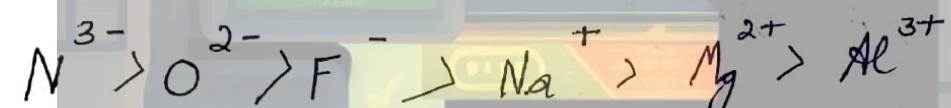
$$[A < A^- < A^{2-} < A^{3-}]$$

Note :- Size

$$A^{3+} < A^{2+} < A^+ < A < A^- < A^{2-} < A^{3-}$$



Size or Radius comparison between iso-electronic species.

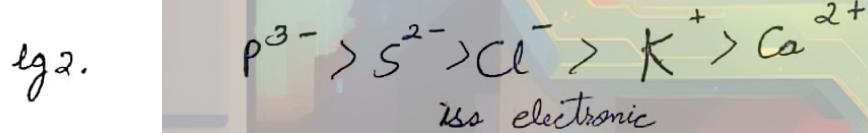


$Z \uparrow$

iso-electronic

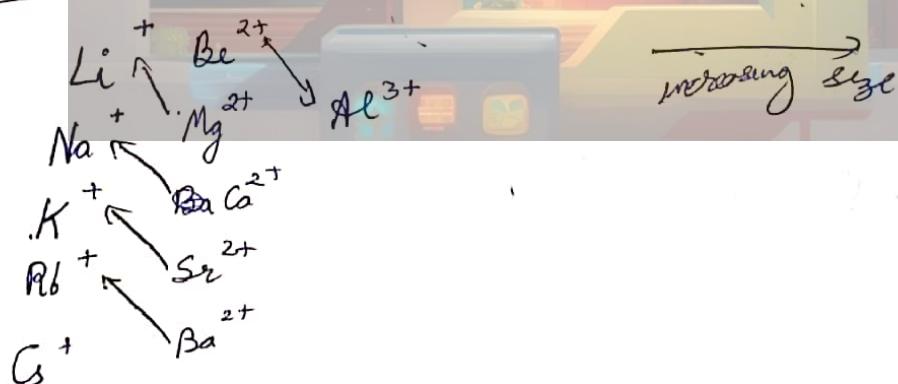
$O = \text{some}$

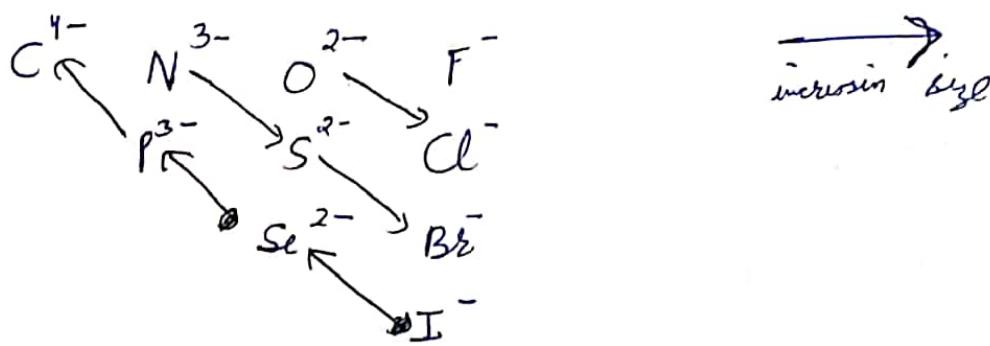
$Z \uparrow \rightarrow r \downarrow$, Size \downarrow



iso electronic

Some Mixed Order.





① Radius Order?

① K, Rb, Cs

~~K < Rb < Cs~~ ✓

② Be, F, O, N, Ne

~~Ne > Be > O > N > F~~ ✓

③ Mg, Na, K, Rb

Mg < Na < K < Rb ✓

④ Sb, P, As, S

S < P < As < Sb ✓

⑤ P, As, Cl, S

Cl < S < P < As ✓

⑥ Sc, V, Mn

Sc > V > Mn ✓

⑦ Fe, Co, Ni

Fe ≈ Co ≈ Ni ✓

⑧ Cu, Zn

Cu < Zn ✓

⑨ Fe, Fe^{2+} , Fe^{3+}

$\text{Fe} > \text{Fe}^{2+} > \text{Fe}^{3+} \checkmark$

⑩ O, O^{-1} , O^{-2}

~~$\text{O}^{-1} > \text{O} > \text{O}^{2+}$~~ $\text{O}^{2-} > \text{O}^{-1} > \text{O}$

⑪ Pd, Ni, Pt

Ni < Pd \approx Pt ✓

⑫ V, Ta, Nb

V < Nb \approx Ta ✓

⑬ Sb, S, Se, Te

S < Se \approx Te < Sb ✓

⑭ Xe, Rn, Kr

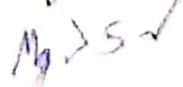
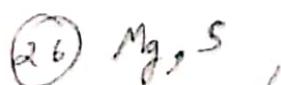
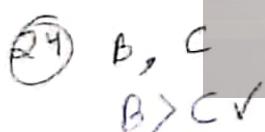
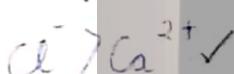
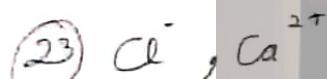
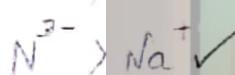
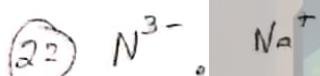
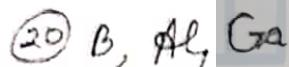
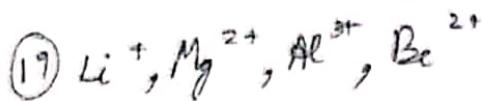
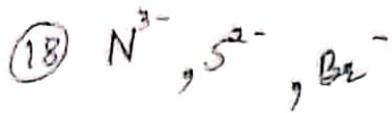
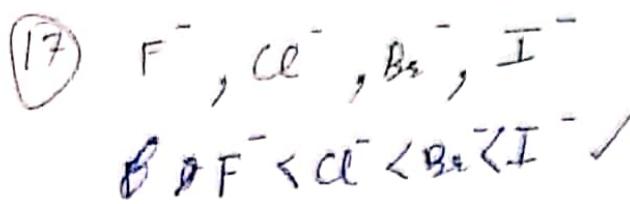
Kr < Xe < Rn ✓

⑮ $\text{A}^+, \text{A}^{2+}, \text{A}^{2-}, \text{A}^-, \text{A}^0$

$\text{A}^{2-} > \text{A}^- > \text{A}^0 > \text{A}^+ > \text{A}^{2+} \checkmark$

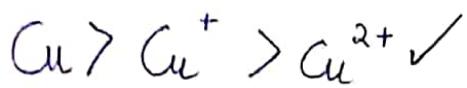
⑯ μ^- , $\text{Li}^{\oplus +}$

$\mu^- > \text{Li}^+ \checkmark$

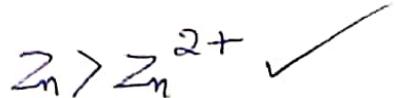


97

(27) $\text{Cu}, \text{Cu}^+, \text{Cu}^{2+}$



(28) $\text{Zn}, \text{Zn}^{2+}$



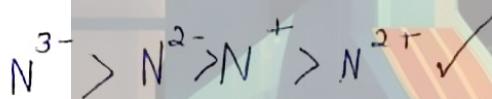
(29) P, As



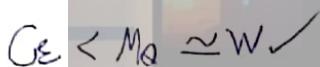
(30) $\alpha X, \text{Kr}$



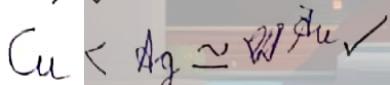
(31) $\text{N}^{3-}, \text{N}^{2-}, \text{N}^+, \text{N}^{2+}$



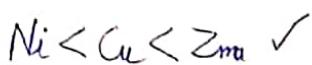
(32) Ge, Mo, W



(33) Cu, Ag, Au



(34) Ni, Zn, Cu

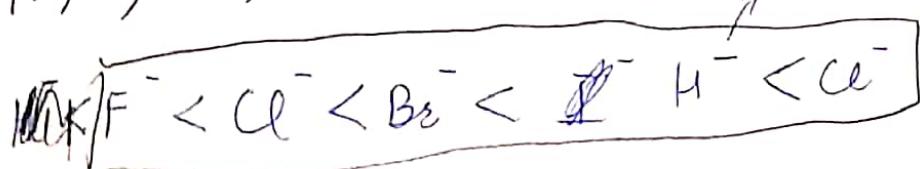


(35) Cl, Cl^+ , Cl^-



* high repulsion
bcs of extra e-

(36) $\text{H}^-, \text{F}^-, \text{Cl}^-, \text{Br}^-, \text{I}^-$



(38)

H.W. 29-04-2024

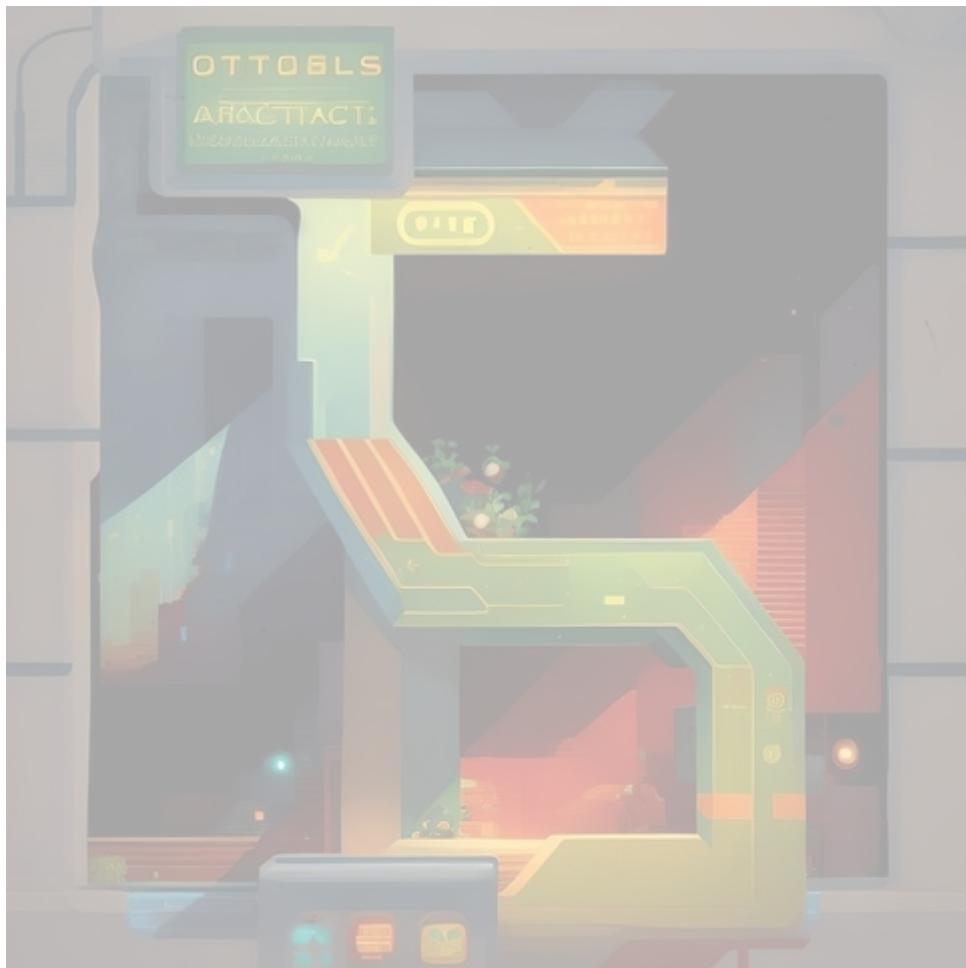
~~O~~. O-1 (Q1-Q14)

O-2 (Q1,2)

S-1 (Q13)

J M (3, 6, 7, 13, 15, 18, 21, 25, 27, 32, 35)

J A (2,6)



Note - Largest Cation - Cs^+

Smallest cation - H^+

Largest mono negative anion = I^-

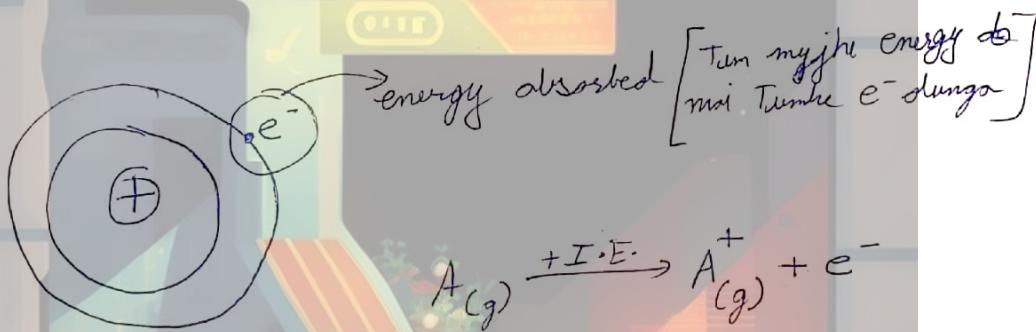
Second - largest mono negative anion = H^-

Smallest mono - negative anion = F^-

* Ionisation Energy / Ionisation Potential (I.P.)

at Constant temp & pressure

→ It is the minimum amount of energy required to remove the outermost electron from isolated gaseous atom or ion.



→ Unit of I.E. :- kJ/mol $\Rightarrow \text{e.v/mol}$, Kcal/mol , J/mol , Col/mol

↓
per electron
volt

Enthalpy - It is the heat content of system.



$$\Delta H (\text{change in heat/enthalpy}) = E_P - E_R$$

$\Delta H = \text{① } \text{Ve} \rightarrow$ Endothermic reaction \rightarrow Energy absorbed

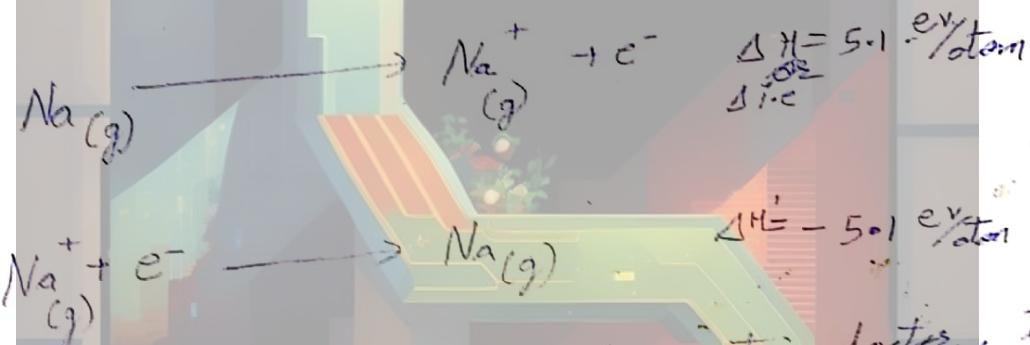
$\Delta H = \text{② } \text{Ve} \rightarrow$ Exothermic Reaction \rightarrow Energy released.

$\Delta H_{I.E}$ (Ionization Enthalpy) :- It is the change in enthalpy when 1 mole of outermost electrons are removed from 1 mole of isolated gaseous atoms/ions.

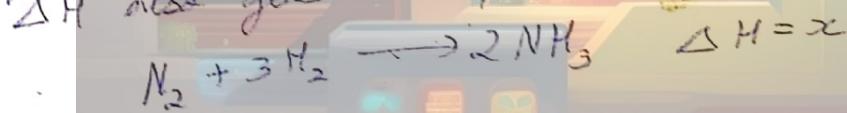
Properties of Enthalpy of reactions :-

Properties of Enthalpy of reactions :-

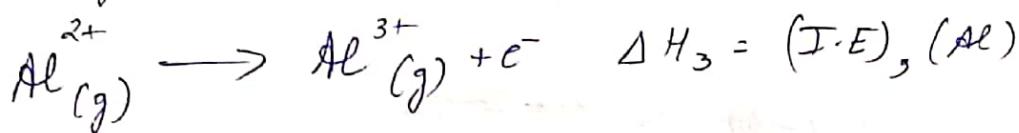
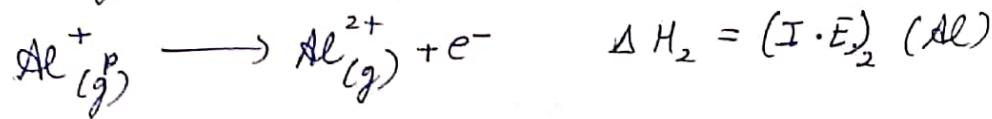
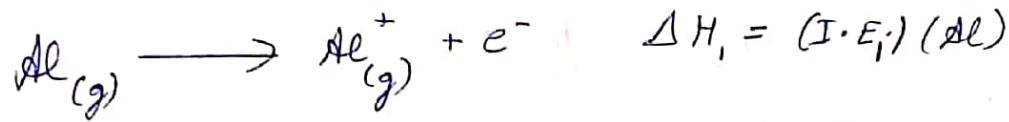
- ① If the reaction is reversed, the sign of ΔH is also reversed.



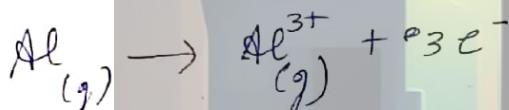
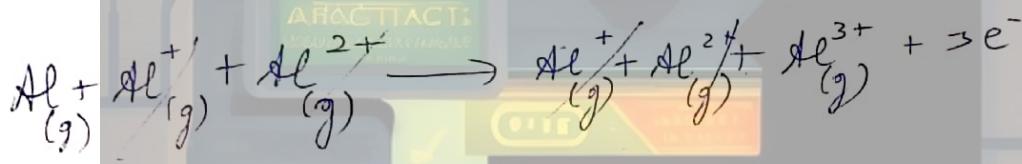
- ② If the reaction is multiplied by certain factor, then ΔH also gets multiplied by that factor.



③ If two or more than two reactions are added, their ΔH are also added

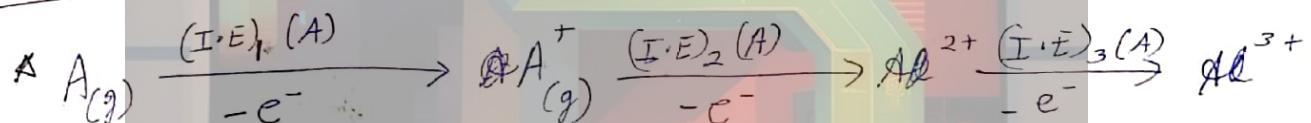


Adding the three



$$\boxed{\Delta H = \Delta H_1 + \Delta H_2 + \Delta H_3}$$

Successive Ionisation Energy

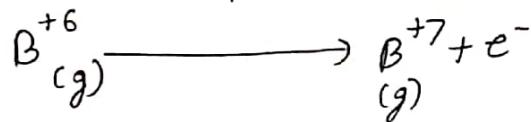


$$(I \cdot E)_1(A) = (I \cdot E)_2(A^+)$$

$$(I \cdot E)_3(A) = (I \cdot E)_2(A^+) = (I \cdot E)_3(A^{2+})$$

Note:- $\boxed{(I \cdot E)_1 < (I \cdot E)_2 < (I \cdot E)_3}$

Q which of the following I.E's correspond to the process



- A) I.E₁(B)
✓ B) I.E₆(B)
✓ C) I.E₇(B)
✓ D) I.E₅(B²⁺) ✓

- ✓ E) I.E₄(B³⁺) ✓
✓ F) I.E₃(B⁴⁺) ✓
✓ G) I.E₁(B⁵⁺) ✓
✓ H) I.E₂(B⁶⁺) ✓
✓ I) I.E₆(B⁺¹) ✓

Q

Valence e⁻ calculation

~~(I.E)~~₁(x)

Q. An element X, find valence shell e⁻

$$(I.E)_1(x) = 5 \text{ ev/atom}$$

$$(I.E)_2(x) = 13 \text{ ev/atom}$$

$$(I.E)_3(x) = 80 \text{ ev/atom}$$

$$(I.E)_4(x) = 100 \text{ ev/atom}$$

valence shell has 2e⁻

Q. Identify X

$$(I.E)_1(x) = 8 \text{ ev/atom}$$

A) Be

$$(I.E)_2(x) = 14 \text{ ev/atom}$$

B) Na

$$(I.E)_3(x) = 25 \text{ ev/atom}$$

C) Al

$$(I.E)_4(x) = 140 \text{ ev/atom}$$

D) B

$$(I.E)_5(x) = 172 \text{ ev/atom}$$

$$(I.E)_6(x) = 210 \text{ ev/atom}$$

Q3. find the no. of valence shell e^- in element A.

$$I.E_1 = 28 \text{ eV/atom}$$

$$I.E_2 = 39 \text{ eV/atom}$$

$$I.E_3 = 64 \text{ eV/atom}$$

$$I.E_4 = 102 \text{ eV/atom}$$

$$I.E.5 = 775 \text{ eV/atom}$$

$$I.E.6 = 1020 \text{ eV/atom}$$

$$\text{Valence Shell } e^- = 3$$

factors affecting I.E. →

① Effective Nuclear Charge (Z_{eff}) :- Ionisation energy is directly proportional to Z_{eff} .

$$Z_{eff} \uparrow \wedge I.E. \uparrow$$

Size ↓

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

② value of n :

$$n \uparrow; \text{size} \uparrow; I.E. \downarrow$$

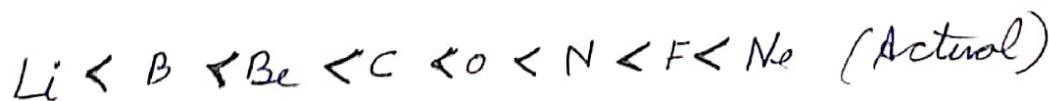
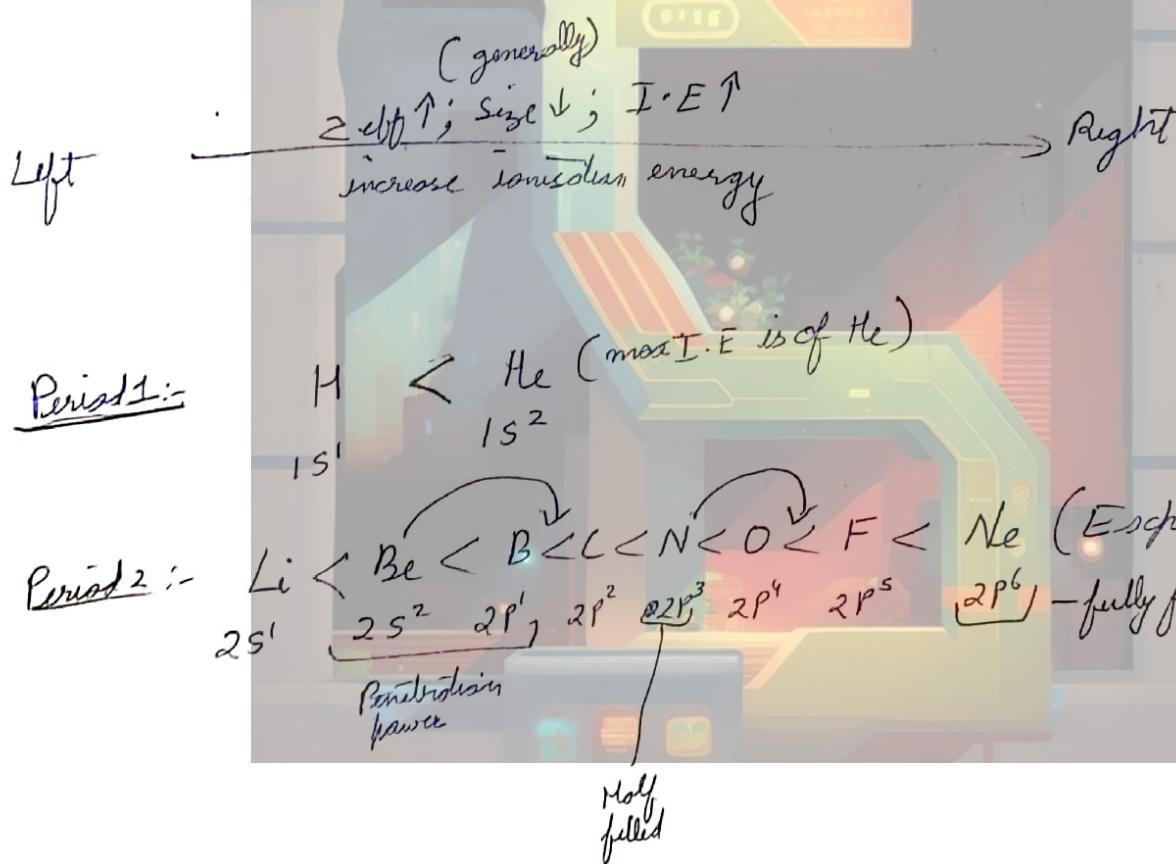
$$I.E. \propto \frac{1}{\text{size}} \propto \frac{1}{n}$$

③ Type of subshell -

$$ns > np > nd > nf$$

④ Half & fully filled e⁻ config:- Half & fully filled e⁻ config are more stable hence process of e⁻ removal will be difficult for such configuration -

Variation along period - Variation of ionization energy across period.



Reason :- ① Due to high penetration of $2s$ subshell compared to $2p$ subshell, the removal of e^- from $2s$ of Be requires more energy as compared to energy required for removal of e^- from $2p$ of boron.

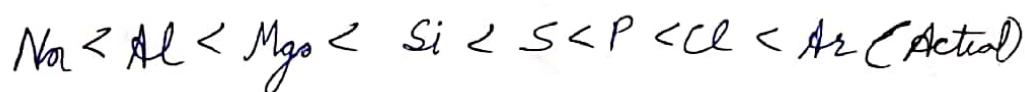
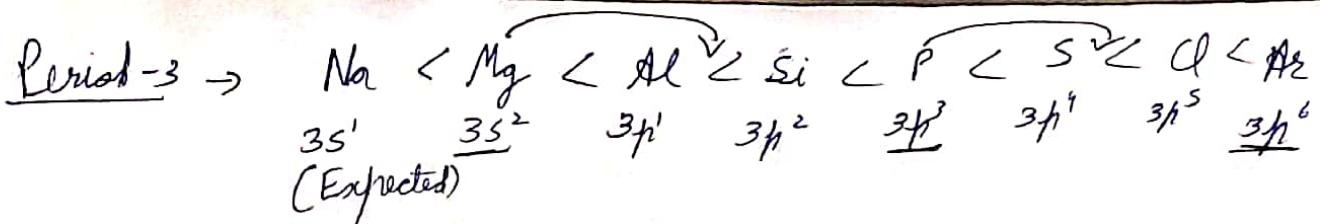
The configuration of Be is $1s^2 2s^2$ which is more stable configuration therefore to remove an e^- from Be its stability should be broken first which requires energy. Also now e^- will be removed from $2s$ which is closer to nucleus thereby requiring more energy.

Hence Overall energy requirement is high.

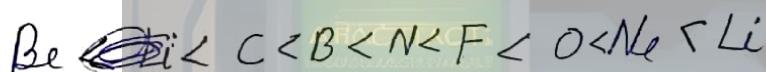
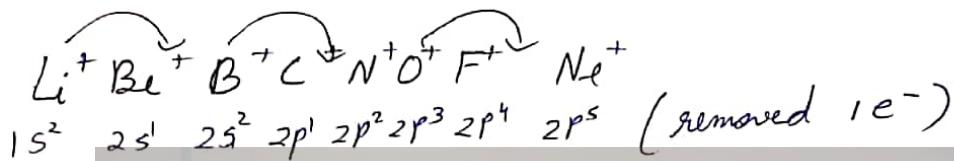
② In case of Nitrogen and oxygen, Nitrogen has half filled stability therefore to remove an electron more energy is required.

$$I.E.(N) > I.E(O)$$

③ Neon has highest ionisation energy in 2nd period due to fully filled electronic configuration and very high Z_{eff} .



Q1. ($I.E_3$) order



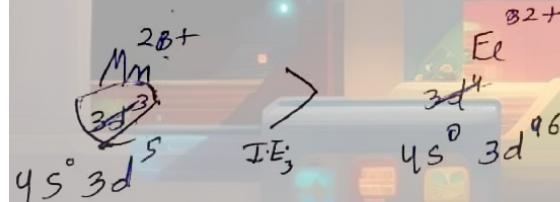
Note - In a periodic table He has highest Ionisation Energy.

② In periodic Table Cesium ~~has~~ has has highest ionisation energy. Cs/Fr

③ In D-block from left to right no regular order was observed hence need not to study

Q2.

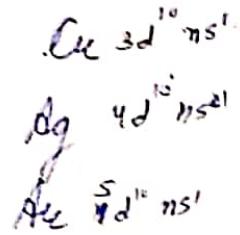
$I.E_3 \quad \text{Mn} \quad I.E_3 \quad \text{Fe}$



↓
Half filled

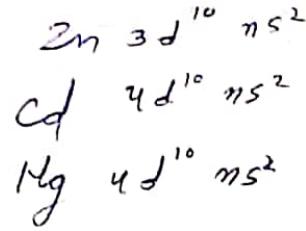
Q3. I.E.

Group 11



$<$
 $<$
 $<$

Group 12



Variation of I.E. down the group:-

Top



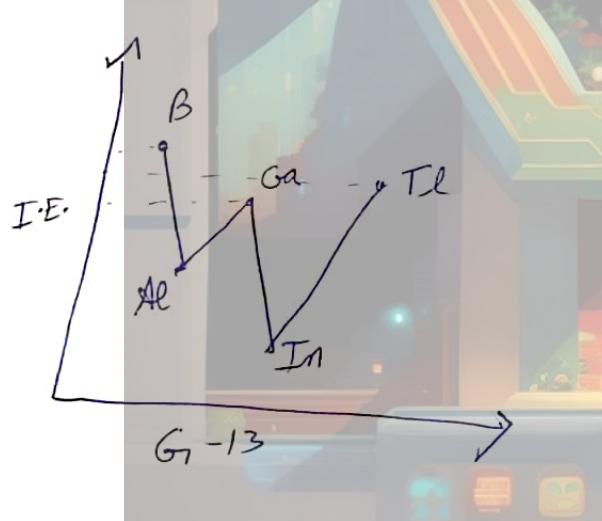
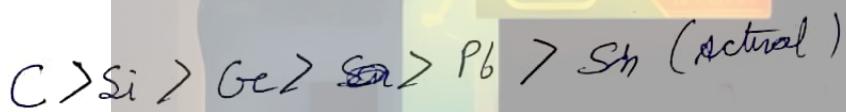
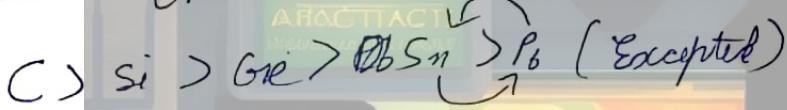
	G 1	G 2	G 13	G 14	G 15	G 16	G 17	G 18
Li	Be	E	C	N	O	F	Ne	
✓	✓	X	✓	✓	✓	✓	✓	
Na	Mg		Si	P	S	Cl	Az	
✓	✓		✓	✓	✓	✓	✓	
K	Ca	E	Ge	As	Se	Br	Kz	
✓	✓	C	✓	✓	✓	✓	✓	
Rb	Si	T	Se	Sb	Te	I	Xe	
✓	✓	T	✓	✓	✓	✓	✓	
Cs	Ba	I	Pb	Bi	P	At	Rn	
		R	V					
			Sn					

Exception 1 - G₇-13



	Actual
B	1
Al	5
Ga	3
In	2
Tl	4

Exception 2 - G₇-14



Explanation (Exception 1) — In Boron family, Due to lanthanides and D-block contraction in Tl, its Z_{eff} increases sharply due to which Ionisation Energy also increases and becomes next to boron.

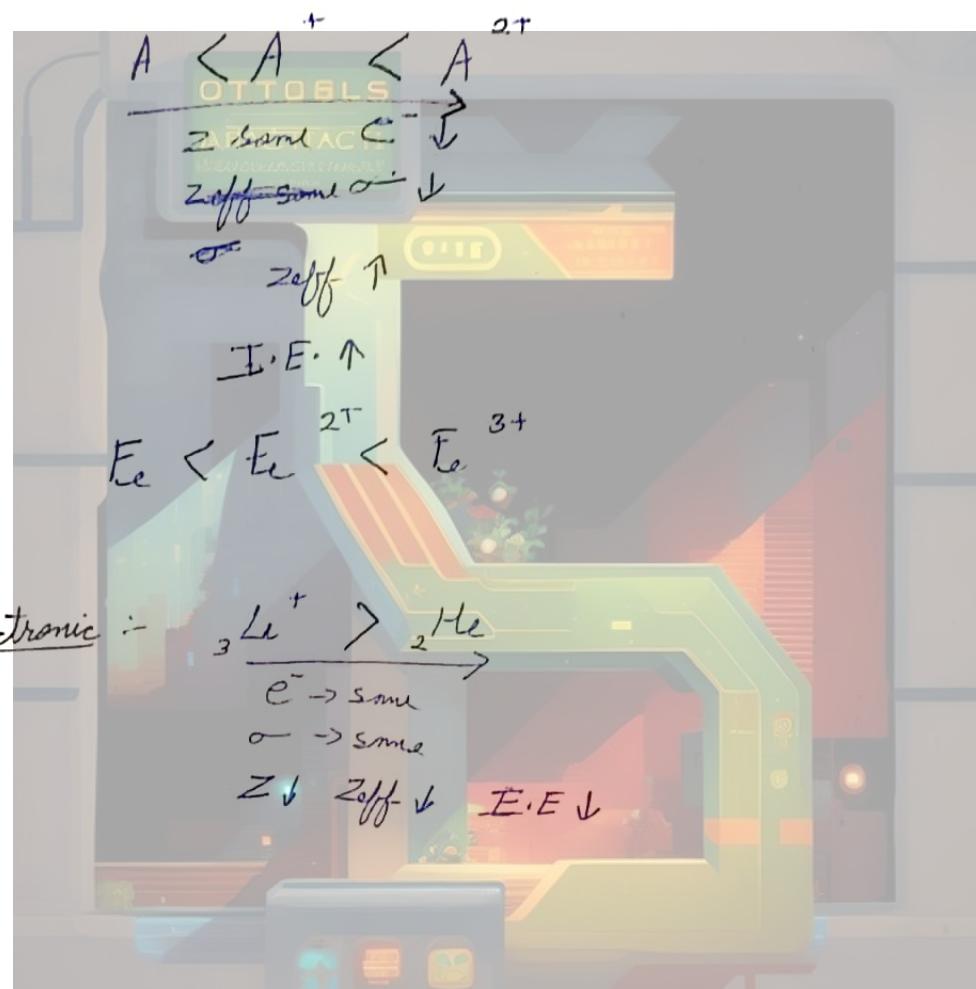
Similar effect is observed with respect to third Lead (Sn, Pb) but ionisation energy of lead does not increase too much like thallium.

D-Block

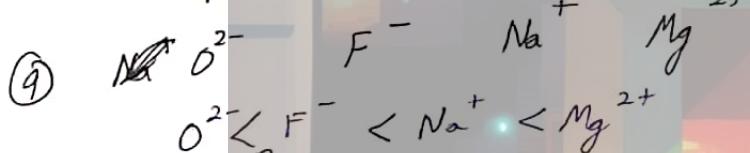
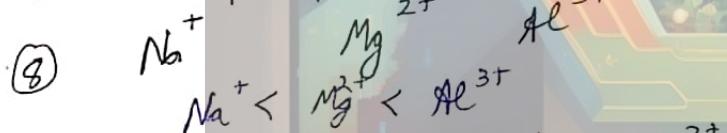
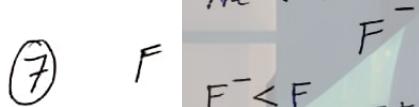
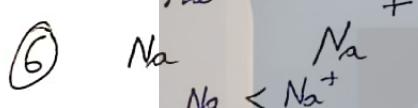
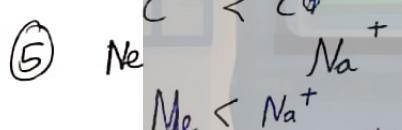
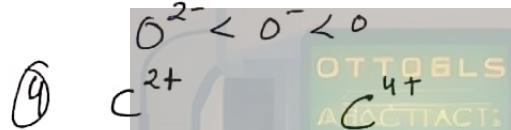
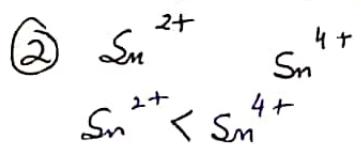
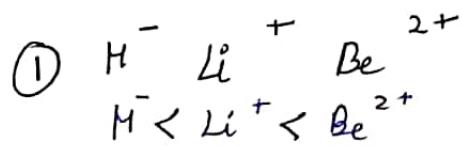
Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn
✓	ss								
Y	Zr	Nb	Mo	Tc	Ru	Rh	Pt	Ag	Cd
✓	VA	VA	VA	X	X	X	X	X	X
La	Hf	Ta	W	Re	Os	Ir	Pd	Au	Pt

I_{so} - Protonic :-

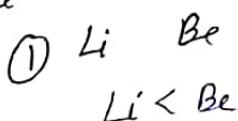
e.g. 1.



Q① I.E order



Q②:



(11)

③ B C N
 $B < C < N$

④ O N F
 $O < N < F$

⑤ N O
 $O < N$

⑥ Na Al Mg Si
 $Na < Al < Mg < Si$

⑦ Cl Br I
~~Cl > B~~ $I < Br < Cl$

⑧ S O F
 $S < O < F$

⑨ He Kr Xe
 $Xe < Kr < He$

⑩ K Na Mg
 $K < Na < Mg$

⑪ P As Sb
 $Sb < As < P$

⑫ B Ga Al
 $B, Al < Ga < B$

⑬ C Si Sn
 $Sn < Si < C$

(14) Ge Mn

$$\text{Ge} < \text{Mn} \quad \checkmark$$

(15) V Nb Ta

$$\text{V} > \text{Nb} \simeq \text{Ta} \quad \checkmark \checkmark$$

(16) Sc Y La

$$\text{Sc} > \text{Y} > \text{La} \quad \checkmark$$

(17) H⁻ Li⁺ Be²⁺

$$\text{H}^- < \text{Li}^+ < \text{Be}^{2+} \quad \checkmark$$

Application of I.E -

① As ionisation energy

I.E.↑, M.C.↓

increases, metallic character decreases

Ex :- Li > Na > K > Rb > Cs [I.E. order]

Metallic character order

$$\text{Li} < \text{Na} < \text{K} < \text{Rb} < \text{Cs}$$

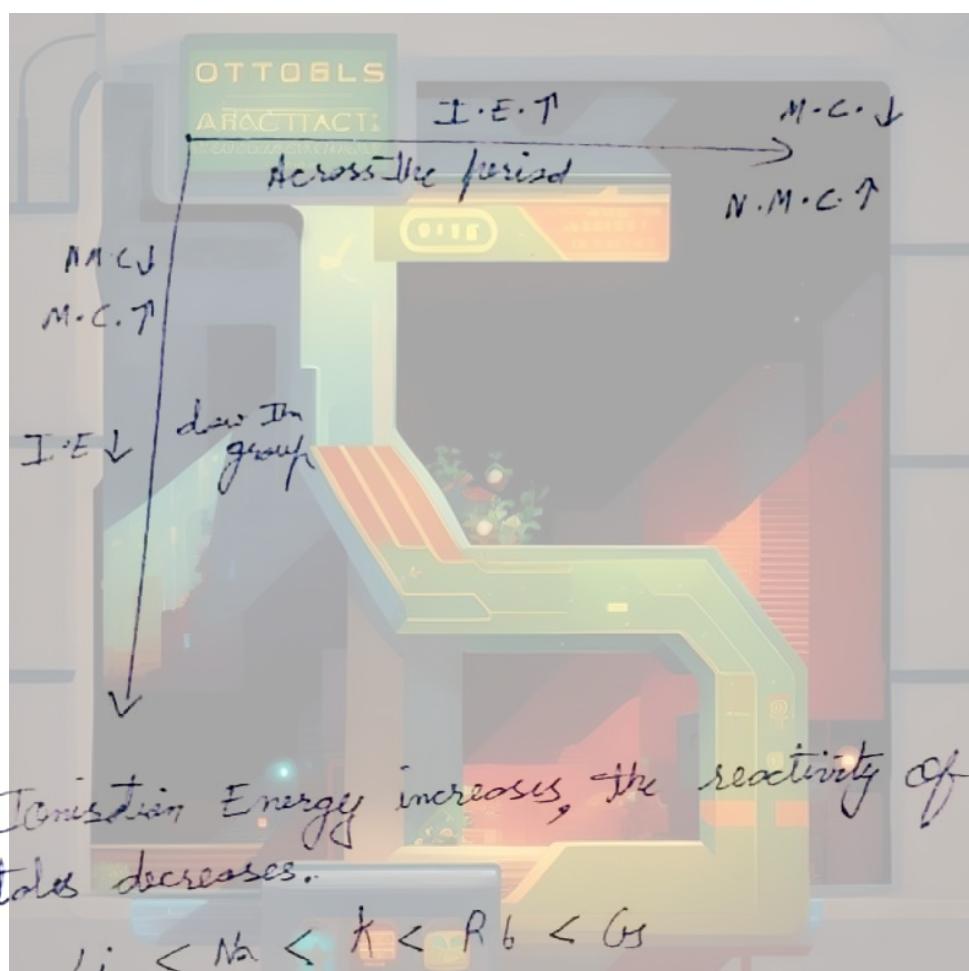
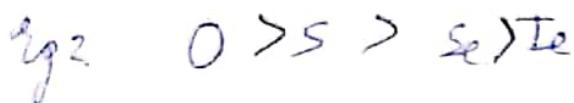
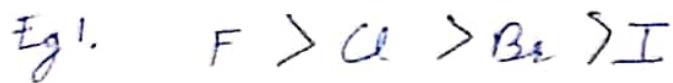
I.E.↑, M.C.↑

Eg 2. $\frac{\text{Be}}{2s} < \frac{\text{B}}{2p}$

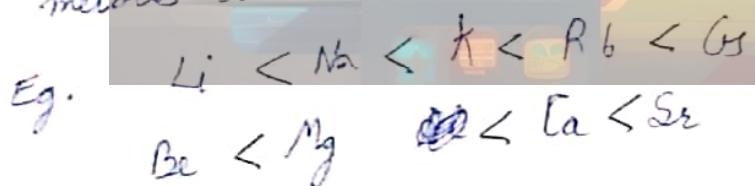
Eg 3. $\text{Be} < \text{Mg}$

② As the ionisation energy increases, non-metallic character increases.

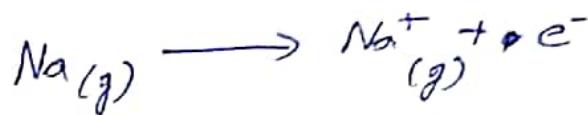
$$I.E \uparrow \quad N.M.C \uparrow$$



③ As Ionisation Energy increases, the reactivity of metals decreases.



④ Oxidation \rightarrow It means removal of e^- .



Reduction \rightarrow Addition of e^- .



Oxidising Power \rightarrow

Oxidising Power \rightarrow It means the tendency of an element to oxidise other elements by reducing itself.

Reducing Power \rightarrow Tendency of an element to reduce other elements by oxidising itself. (Oxidise easily \Rightarrow loses electrons)

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

$$\underline{I.E. \downarrow \quad \text{Reducing Power} \uparrow}$$

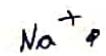
⑤ ② If

$$(I.E.)_{n+1} - (I.E.)_n > 16 \text{ eV/atom}$$

Then lower oxidation state (n) will be more stable.

Na

eV/atom



IE₁

05.1



IE₂

41.2

36.1 Na⁺ more stable

IE₃

71.62

b) $(I \cdot E)_{n+1} - (I \cdot E)_n < 11 \text{ eV/atom}$

then higher oxidation will be more stable

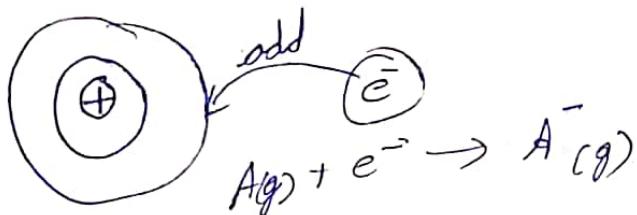
Eg2. $(I \cdot E)_2 - (I \cdot E)_1 = 7.4 \text{ eV/atom} < 11 \text{ eV/atom}$

(Mg)

Mg²⁺ is more stable than Mg⁺

116

Electron Affinity (E.A.) - amount of energy released when an electron is added to the valence shell of an isolated gaseous atom or ion.

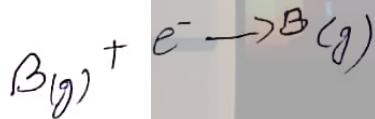


let energy released = 100 kJ/mol

$$E.A = 100 \text{ kJ/mol}$$

$$\Delta H_{E.g} (\text{electron gain enthalpy}) = -100 \text{ kJ/mol} \quad (\text{Exothermic reaction})$$

Electron Gain Enthalpy ($\Delta H_{E.G.}$) - Change in enthalpy when ~~1 mol~~ mole of electrons are added to the outermost shell of 1 mole isolated gaseous atoms.



energy released = 0 (atom ka e⁻ ke liye koi affection nahi)

energy supplied = 100 kJ/mol

$$\Delta H_{E.g} = +100 \text{ kJ/mol}$$

endothermic

(e⁻ zabolardosti olaya)

at 0K (0 Kelvin)

$$\Delta H_{E.G.} = -E.A$$

Note:- ① $\Delta H_{Eg} = -\text{ve}$; $E.A = +\text{ve} \Rightarrow$ energy released

② $\Delta H_{Eg} = +\text{ve}$; $E.A = -\text{ve}$ or zero \Rightarrow energy absorbed

Note:-

① In case of inert gas or some stable configuration energy is ~~does~~ not released, instead of ~~this~~, we have to supply energy.

Ex Be, Mg, N and all inert gases (He, Ne, Ar, Kr, Xe, Rn) follow this and absorb energy on addition of electron.

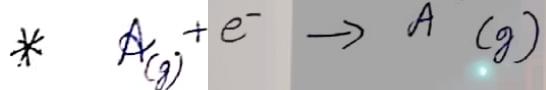
$$\Delta H_{Eg_1} = +\text{ve}$$

$$E.A_1 = -\text{ve}$$

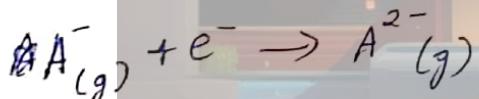
Rest all atoms except Be, Mg, N and inert gases

$$\Delta H_{Eg_1} = -\text{ve}$$

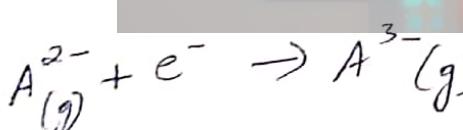
$$E.A_1 = +\text{ve}$$



$$\Delta H_{Eg_1} \text{ or } E.A_1$$



$$\Delta H_{Eg_2} \text{ or } E.A_2$$



$$\Delta H_{Eg_3} \text{ or } E.A_3$$

For Any Atom $\rightarrow \boxed{\Delta H_{Eg_1} = +\text{ve} (\text{always})}$

Reason:- Since Anion is always negatively charged so, incoming electron faces repulsion from the Anion and its repulsion is dominating.

$$\Delta H \Rightarrow \text{supplied}$$

$$\left[\Delta H_{E.G.E_3} > \Delta H_{E.G.E_2} > \Delta H_{E.G.E_1} \right]$$

\oplus ve

\ominus ve or \oplus ve

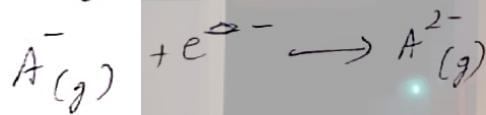
Note- $E_{A_1^-}$ can be positive or negative

($\text{Be}, \text{Mg}, \text{N}$, All inert gases)

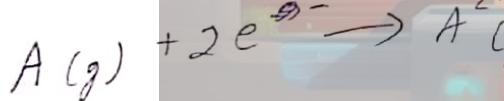
$\ominus \Delta H_{E.G.E_1}$ can be \ominus ve or \oplus ve



$\Delta H_{E.G.E_1} \Rightarrow \oplus \text{ve} / \ominus \text{ve}$



$\Delta H_{E.G.E_2} \Rightarrow \oplus \text{ve always}$

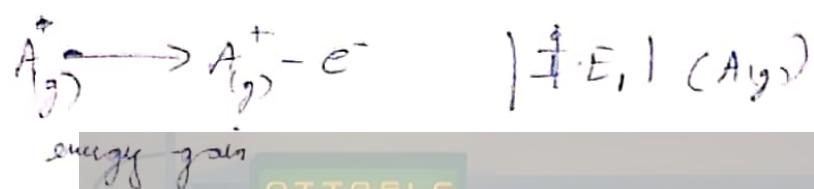
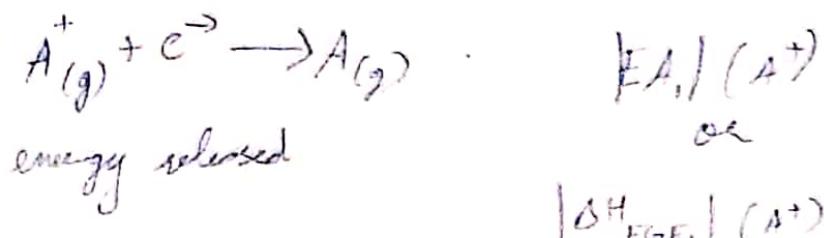


$$\Delta H_{E.G.E_{\text{Total}}} = \Delta H_{E.G.E_1} + \Delta H_{E.G.E_2}$$

$\ominus \text{ve} / \oplus \text{ve} + \oplus \text{ve}$

$\oplus \text{ve always}$

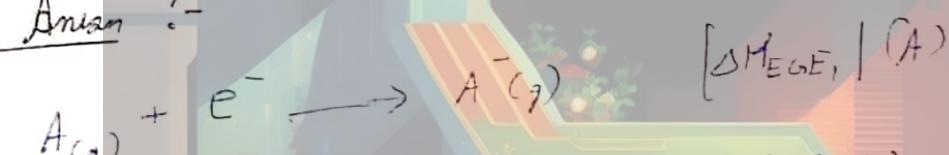
For Cations :-



$$| \Delta H_{E.G.E.1}(A^+) = | I.E.1(A)$$

$$| EA_1(A^+) = | I.E.1(A)$$

For Anion :-



$$| \Delta H_{E.G.E.1}(A_{(g)}) = | I.E.1(A_{(g)})$$

~~Trends in E.G.~~

Trends in E.G.

① $Z_{\text{eff}} \uparrow$ Attract \uparrow E.A. \uparrow

② value of n : $n \uparrow$; size \uparrow ; $Z_{\text{eff}} \downarrow$; E.A. \downarrow

③ Type of Subshell in which e^- is added - (Penetration power)

EA: $ns > np > nd > nf$
के निकट से प्रवाह करते हैं इनस की अवधि ज्यादा होती है।

④ Half filled & fully filled - for such atoms E.A. is very less or zero in magnitude.

In Periodic Table

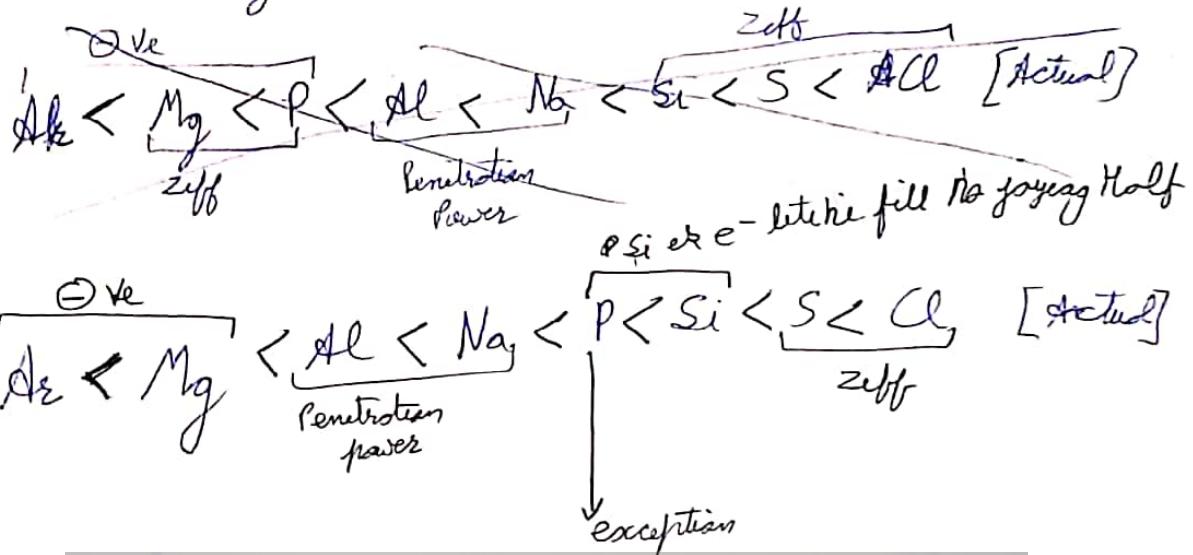
$Z_{\text{eff}} \uparrow$; E.A. \uparrow (in general)

Period - 2 $Li < Be < B < C < N < O < F < Ne$ (Expected)

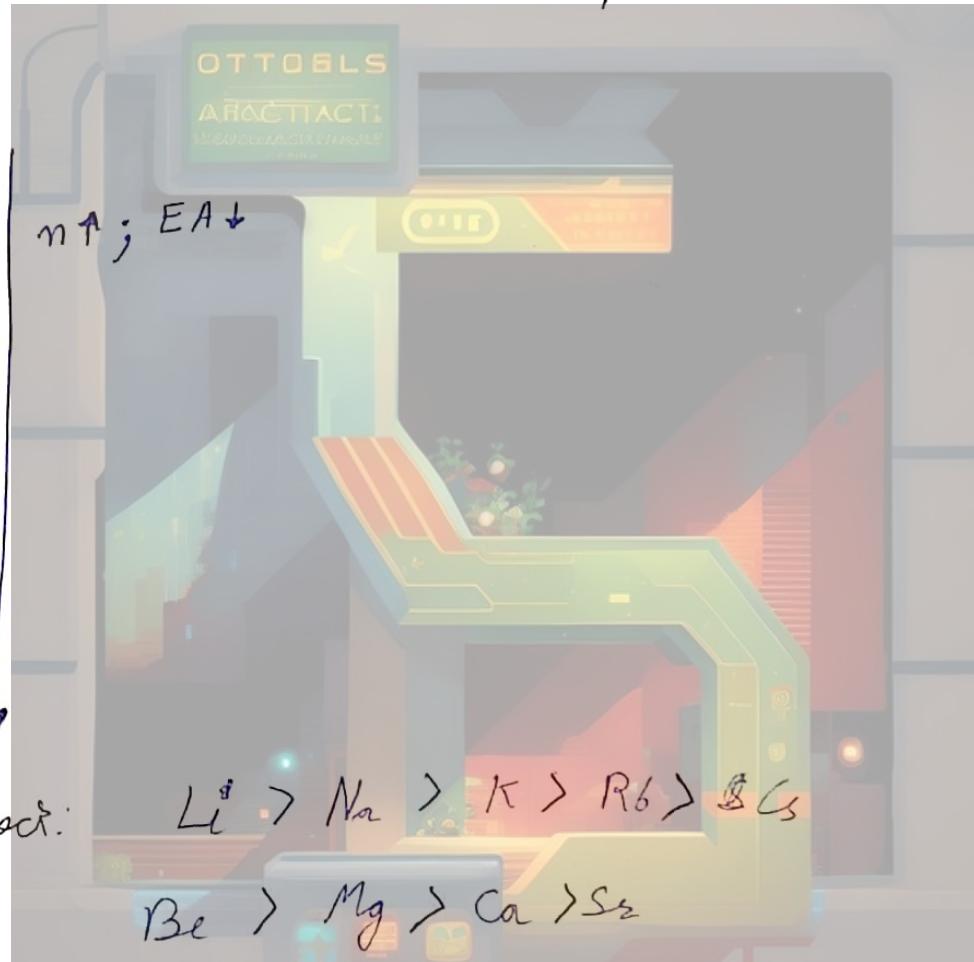
$Ne < Be < N < B < Li < C < O < F$ (Expected)
 Z_{eff} Penetration power

Period - 1 $H > He$

Period-3 $\text{Na} < \text{Mg} < \text{Al} < \text{Si} < \text{P} < \text{S} < \text{Cl} < \text{Ar}$ [Expected]

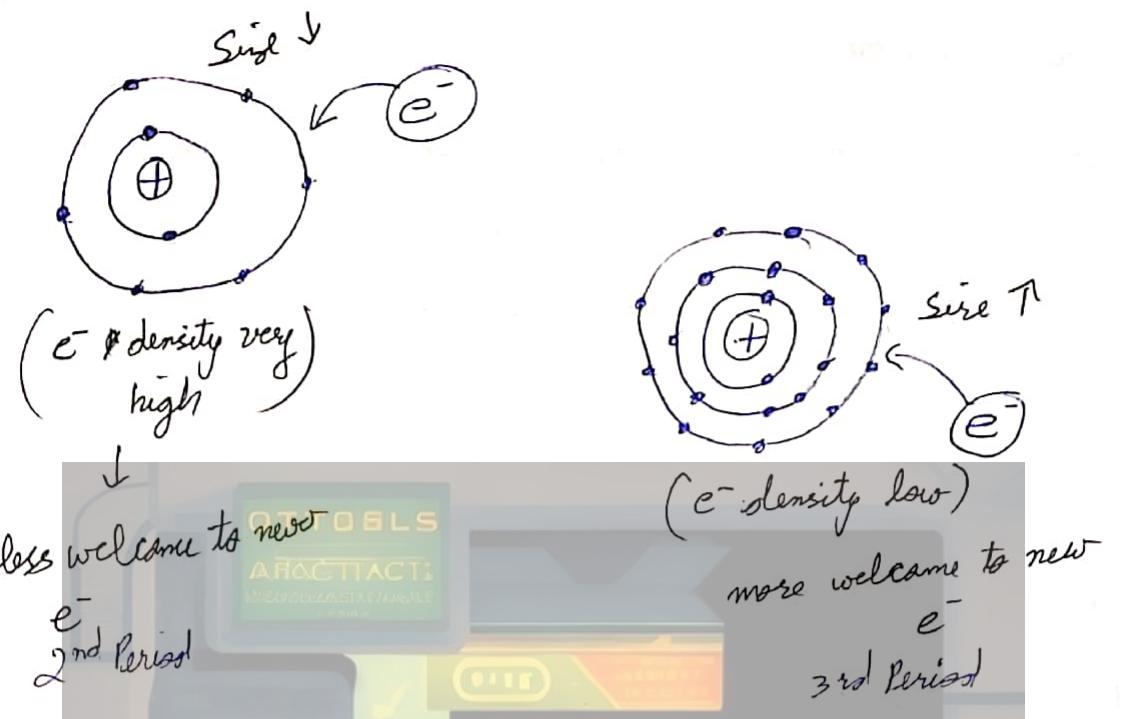


Top



2-nd Period :- $\text{B} \wedge \text{C} \wedge \text{N} \wedge \text{O} \wedge \text{F}$
 3-nd Period :- $\text{Al} \wedge \text{Si} \wedge \text{P} \wedge \text{S} \wedge \text{Cl}$

Reason:-



Due to small size of 2nd Period, P-Block elements, ~~other~~ other e^- which upcoming e^- feels more repulsion cause of other e^- which are already present in the atom.

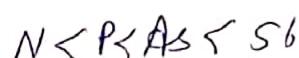
Halogen Family :-



Oxygen Family :-



Nitrogen :-



} experimental based on E.A. Data
(No reason)

Note:-

- ① Oxygen and Nitrogen shows least e⁻ affinity in their respective groups.
 - ② Chlorine has highest e⁻ affinity in periodic table and fluorine comes to 2nd position.
 - ③ As E.A. ↑ ; Non metallic character ↑
- Q1. - E A. order?

① Cl, Br, I

$$I < Br < Cl \checkmark$$

② Na, Al, Mg, Si

$$Mg < Al < Na < Si \checkmark$$

③ F, Cl

$$F < Cl \checkmark$$

④ O, S

$$O < S \checkmark$$

⑤ C, N

$$N < C \checkmark$$

⑥ Se, S

$$Se < S \checkmark$$

⑦ P, As, Sb

$$P < As < Sb \checkmark$$

⑧ Li, Be, B

$$Be < B < Li \checkmark$$

⑨ O, F,
 $O < F \checkmark$

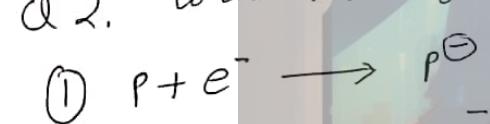
⑩ O, F, Cl
 $O < F < Cl \checkmark$

⑪ N, O, F
 $N < \cancel{O} \quad O < F \checkmark$

⑫ C, O
 $C \cancel{O} < O \checkmark$

⑬ O, N
 $N < O \checkmark$

Q2. Write the sign of ΔH in the following.



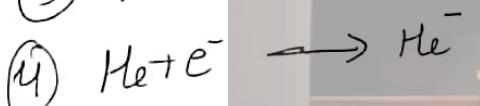
$$\Delta H = \Theta Ve \checkmark$$



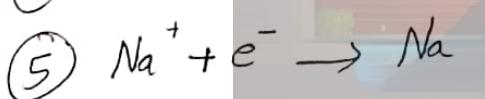
$$\Delta H = \cancel{\Theta} \quad \cancel{\Theta} \quad \Theta Ve \checkmark$$



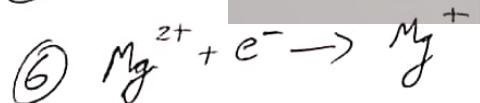
$$\Delta H = \oplus Ve \checkmark$$



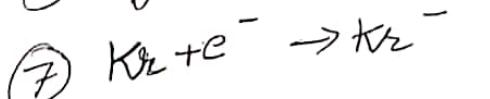
$$\Delta H = \oplus Ve \checkmark$$



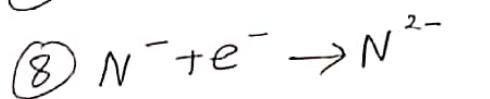
$$\Delta H = \Theta Ve \checkmark$$



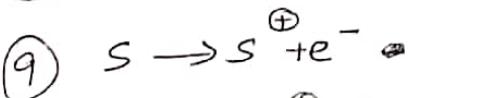
$$\Delta H = \Theta Ve \checkmark$$



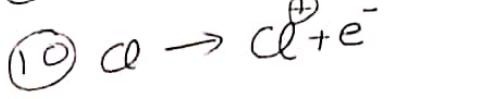
$$\Delta H = \cancel{\Theta} \quad \oplus Ve \checkmark$$



$$\Delta H = \oplus Ve \checkmark$$



$$\Delta H = \Theta Ve \checkmark$$



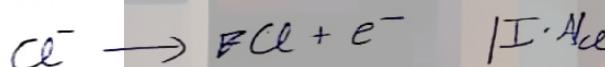
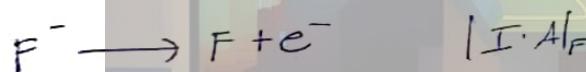
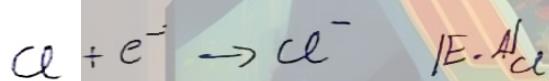
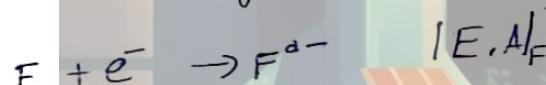
$$\Delta H = \Theta Ve \checkmark$$

Application of E.A.

E.A means \rightarrow Gain of e^- easily \rightarrow Reduce easily
 \downarrow
Oxidise rapidly
 \downarrow
Oxidising Power \uparrow

\rightarrow As The value of E.A. increases, the oxidising Power of the element increases

Q Compare I.E. of F^0 and Cl^-

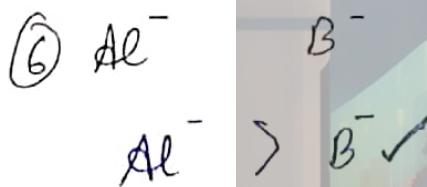
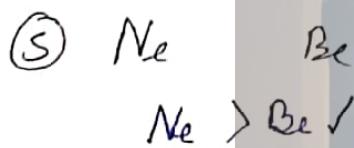
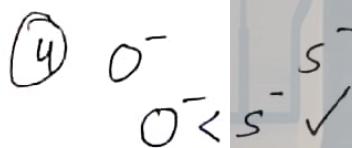
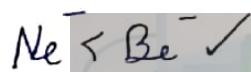
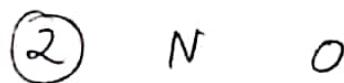
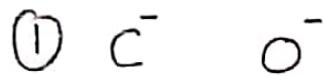


$$E.A = I.A$$

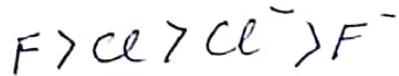
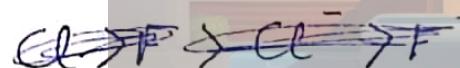
$$E.A_{Cl^-} > E.A_{F^-}$$

$$I.E_{F^-} < I.E_{Cl^-}$$

Q2. Compare I.E. order in following.



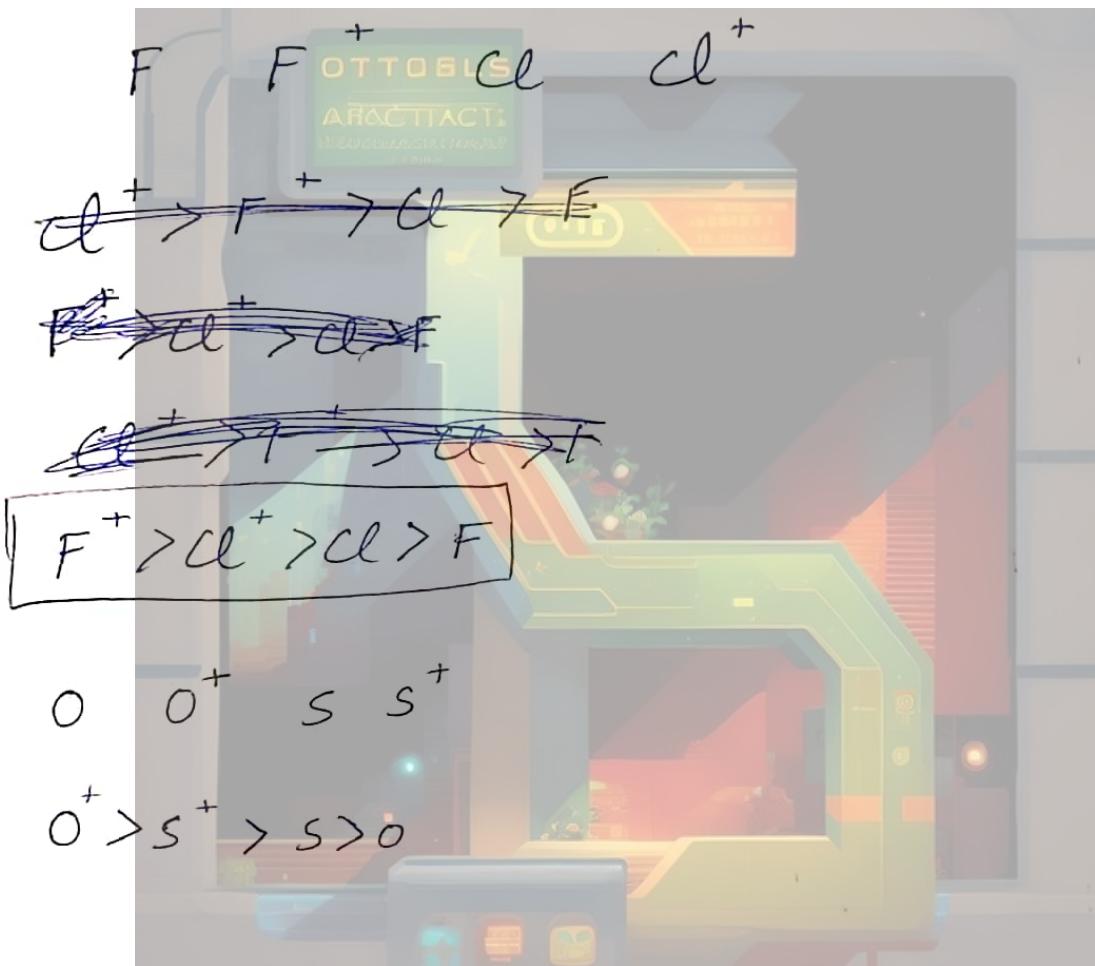
Q3. Compare I.E.



Q 34. IE.



Q S.E. A. order. ?



H.W. 06 - 05 - 2024

O-1 (Q 15-27)

O-2 (Q 3, 4, 5, 6)

~~O-3~~ S-1 (1, 3, 4, 5, 6, 9, 10, 11)

(~~6, 7~~) (~~10, 5~~)

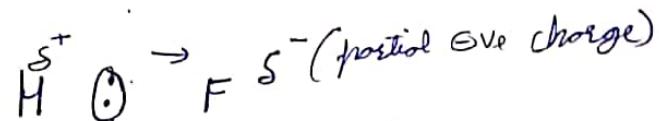
J-A A (Q 3, 1, 4, 5, 8)

J-M (Q 1, 4, 5, 8, 10, 11, 14, 1@.9, 21, 26, 26, 29, -33, 36, 38)

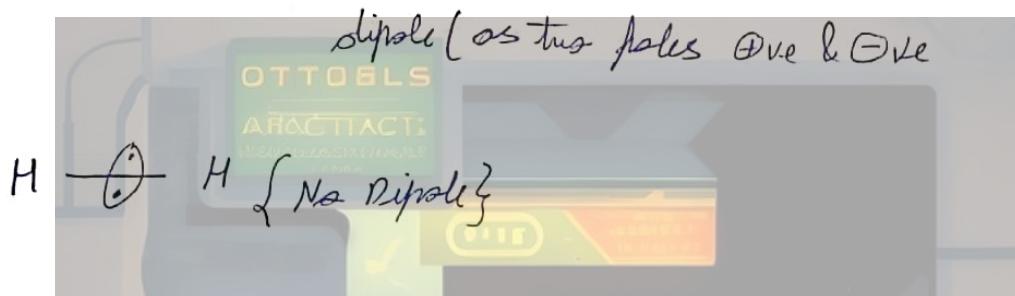


Electron Negativity (E.N) (x)

→ It is the tendency to attract shared pair of electrons by an element in a covalent bonded molecule.



↓ shared pair of e^- moves slightly to F as it is powerful.



→ Electron Negativity has no unit.

→ Three Scales to measure → Pauling Scale, Mulliken Scale, Alfred Rochow Scale.

① Pauling Scale of E.N. :-

→ It is a relative scale of Electron Negativity (Reference element - Fluorine (F)).

→ Used Bond energy data to find Electron Negativity -
for two atoms A & B

$$\Delta E_N = \frac{E_B - E_A}{E_{A-B}} = 0.208 \sqrt{E_{A-B}^{-1} \cdot E_{A-A} \times E_{B-B}}$$

$EN_A \Rightarrow$ Electron Negativity of A

$EN_B \Rightarrow$ Electron Negativity of B

$E_{A-B} \Rightarrow$ Bond energy of A-B bond (~~kJ~~ Kcal/mol)

$E_{A-A} \Rightarrow$ Bond energy of A-A bond (Kcal/mol)

$E_{B-B} \Rightarrow$ Bond energy of B-B bond (Kcal/mol)

$$\Delta EN = |EN_B - EN_A| = 0.102 \sqrt{E_{A+B}} \sqrt{E_{A-A} \times E_{B-B}}$$

~~for kJ/mol~~ for Kilo J/mol

Q1. If $EN_A = 2$

$EN_B > EN_A$

$$E_{A-B} = 76 \text{ Kcal/mol}$$

$$E_{A-A} = 81 \text{ Kcal/mol}$$

$$E_{B-B} = 64 \text{ Kcal/mol}$$

$$EN_B = ?$$

$$EN_B = \cancel{0.2} + 0.208 \sqrt{76 - \sqrt{81 \times 64}}$$
$$= 2 + 0.208 \sqrt{76 - 72}$$
$$= 2 + 0.208 \times 2$$
$$= 2 + 0.416$$

$$EN_B = 2.416$$

Q2. Calculate EN_A if

$$EN_B = 3$$

$$E_{AB} = 65 \text{ kcal/mol}$$

$$E_{A-A} = 49 \text{ kcal/mol}$$

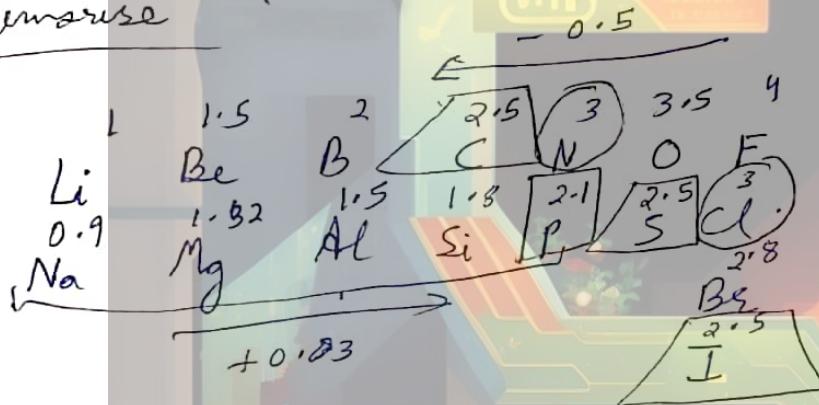
$$E_{B-B} = 64 \text{ kcal/mol}$$

$$EN_A = 3 + 0.208 \times 3$$

$$= 3 + 0.624$$

$$\boxed{= 3.624}$$

Memorise



$$\boxed{H \rightarrow 2.1}$$

$$C \approx I \approx S = 2.5$$

$$H \approx P = 2.1$$

$$N \approx Cl = 3$$

$$F = 4 \text{ (Max)}$$

② Mulliken Scale

$$E.N(x) = \frac{|I.E.(x)| + |E.A.(x)|}{2}$$

I.E, E.N in eV/atom

$$(E.N)_p = \frac{(E.N)_m}{2.8}$$

Q Calculate E.N. (Cl) in Paulin Scale

$$I.E(Cl^-) = 4 \text{ ev/atom} = [E.A]_{Cl}$$

$$E.A(Cl^+) = 13 \text{ ev/atom} = [E.I]_{Cl}$$

$$EN.(m) = \frac{13 + 4}{2}$$

$$= \frac{17}{2}$$

$$(E.N.)_p = \frac{17}{2} \times \frac{1}{2.8}$$

$$= \frac{17}{5.6} = 3.03$$

③ Alfred Roschow Scale :-

$$E.N._{(AR)}(x) = \frac{0.359 Z_{\text{eff}}}{r^2}$$

Z_{eff} → effective nuclear charge

r → covalent radius in \AA°

$$1 \text{ \AA}^\circ = 10^{-10} \text{ m}$$

$$1 \text{ pm} = 10^{-12} \text{ m}$$

$$E.N._P(x) = E.N._{AR} + 0.744$$

$Z_{\text{eff}} \uparrow$; $E.N. \uparrow$

Factors Affecting Electron Negativity

① Z_{eff} :-

$Z_{\text{eff}} \uparrow$

$E.N. \uparrow$

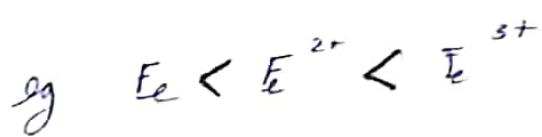
② Value of n :-

$n \uparrow$

$E.N. \downarrow$

③ charge :- (+ve)

$E.N \propto$ oxidation state of element.

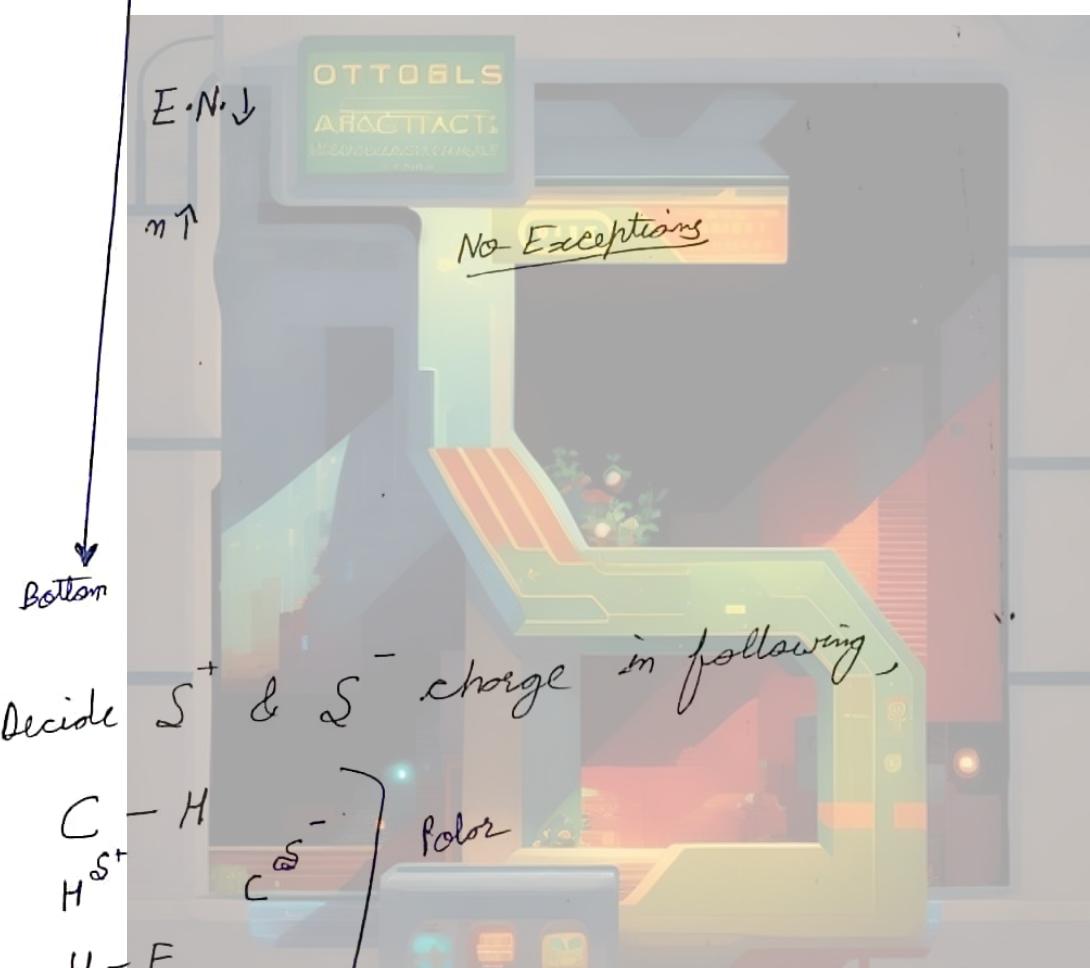


~~Note~~

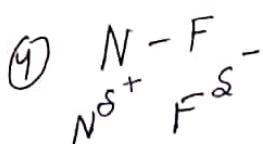
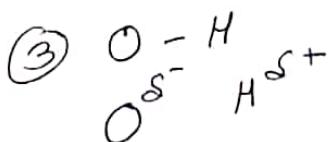
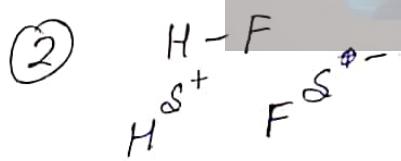
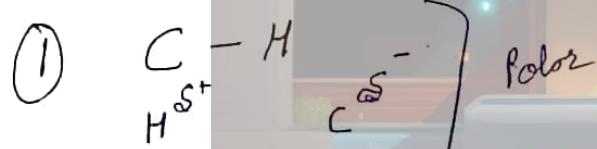
Note - ① No Effect of half filled or fully filled e⁻ config.

Variation across Periodic Table

Left Top $Z_{eff} \uparrow$ E.N.T Right



Q: Decide S^+ & S^- charge in following,



(2)

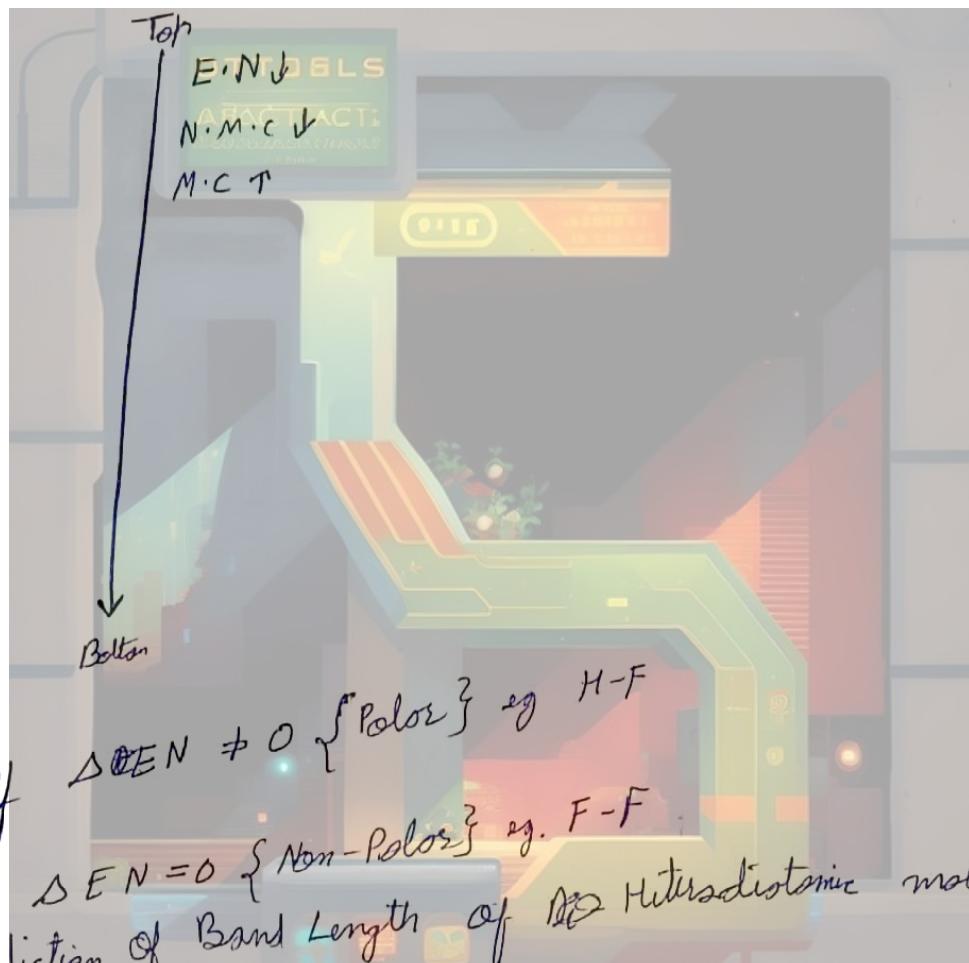
F - F

(5) Non-Polar

Applications of Electron Negativity

→ (1)

$$\xrightarrow[\text{N.C.} \downarrow; \text{N.M.C.} \uparrow]{\text{L. size} \downarrow; \text{Z}_{\text{eff}} \uparrow; \text{EN}^{\uparrow}; \text{I.E}^{\uparrow}; \text{E.A}^{\uparrow}} \text{R}$$



- (2) If $\Delta EN \neq 0$ {Polar} e.g. H-F
 $\Delta EN = 0$ {Non-Polar} e.g. F-F
- (3) Prediction of Bond Length of Heteroditomic molecule.

$$d_{AB} = r_A + r_B + 0.09 |EN_A - EN_B| \text{ (in } \text{\AA})$$

(4) Predicting of % of ionic character in covalent bond.

$$\boxed{\% \text{ Ionic character} = 1.68 / (\Delta EN) + 3.5 (\Delta EN)^2}$$

ΔEN in Pauling Scale

$EN \uparrow ; \% \text{ I.C.} \uparrow ; \% \text{ covalent character} \downarrow$

$$\% \text{ I.C.} + \% \text{ covalent character} = 100$$

Q1. calculate % Ionic character order:

H-F	H-Cl	H-Br	H-I
1.9	0.9	0.7	0.4

$$H-F > H-Cl > H-Br > H-I$$

Q2. Calculate Covalent Character (CC) order in Cl-F bond.

$$EN_{\alpha} = 3$$

$$EN_{N.F} = 4$$

$$\Delta EN = 4 - 3 = 1$$

$$IAC = 16 + 3.5$$

$$\boxed{IAC = 19.5}$$

$$CC = 100 - IAC$$

$$\boxed{CC = 80.5}$$

Q3. I.C% in HCl

$$\Delta EN = 2.1 - 3 \\ = 0.9$$

$$I \cdot C = 16(0.9) + (0.9)^2 \\ = 14.4 + 0.81$$

$$\boxed{= 22.5} \quad = 14.4 + 2.8 \\ = 17.2$$

$$\begin{array}{r} 9 \\ 350 \\ 81 \\ \hline 350 \\ 28000 \\ \hline 28350 \end{array}$$

Q4. ΔEN for $I \cdot C = 50$

$$\Delta EN = x$$

OTTOBLS
ARACTACTIC

$$50 = 16x + \frac{0.81(x^2)}{3.5}$$
$$7x^2 + 10x - 100$$
$$x = -14 \pm \sqrt{196 + 2800}$$
$$= -14 \pm \sqrt{2996}$$
$$= -14 \pm 54$$
$$x = -32 \pm \sqrt{1024 + 2800}$$
$$= -32 \pm \sqrt{3824}$$
$$= -32 \pm 61$$
$$= \frac{-16 \pm 30.5}{7}$$
$$= \frac{14.5}{7}$$
$$\boxed{= 2.1}$$

496
416
104
224
122

Note:- In Real, no bond is 100% ionic or 100% covalent.

$$\Delta EN = 2.1$$

$$\% \text{ I.C} = 50\%$$

$$\% \text{ C.C} = 50\%$$

Predominantly ionic compound

Nature of Compound \rightarrow Ionic

$$\Delta EN < 2.1$$

$$\% \text{ I.C} < 50\%, \% \text{ C.C} > 50\%$$

Covalent compound

$$\Delta EN > 2.1$$

$$\% \text{ I.C} > 50\%, \% \text{ C.C} < 50\%$$

Ionic compound

Q1. If the value of $\sqrt{E_{A-B} - \frac{1}{2}(E_{AA} + E_{BB})}$ is termed as Δ , if Δ is given as 14.4 for elements A and B, and bond length of A-A bond is 1.4 \AA and bond length of B-B bond is 1.2 \AA . Calculate -

- Calculate I.C of A-B bond.
- Covalent bond length of A-B bond.

It can
be written as $K \text{ kJ/mol}$

$$\Delta EN = 0.102 \times 14.4$$

$$= 1.4688$$

$$\begin{array}{r} 144 \\ 102 \\ \hline 288 \end{array}$$

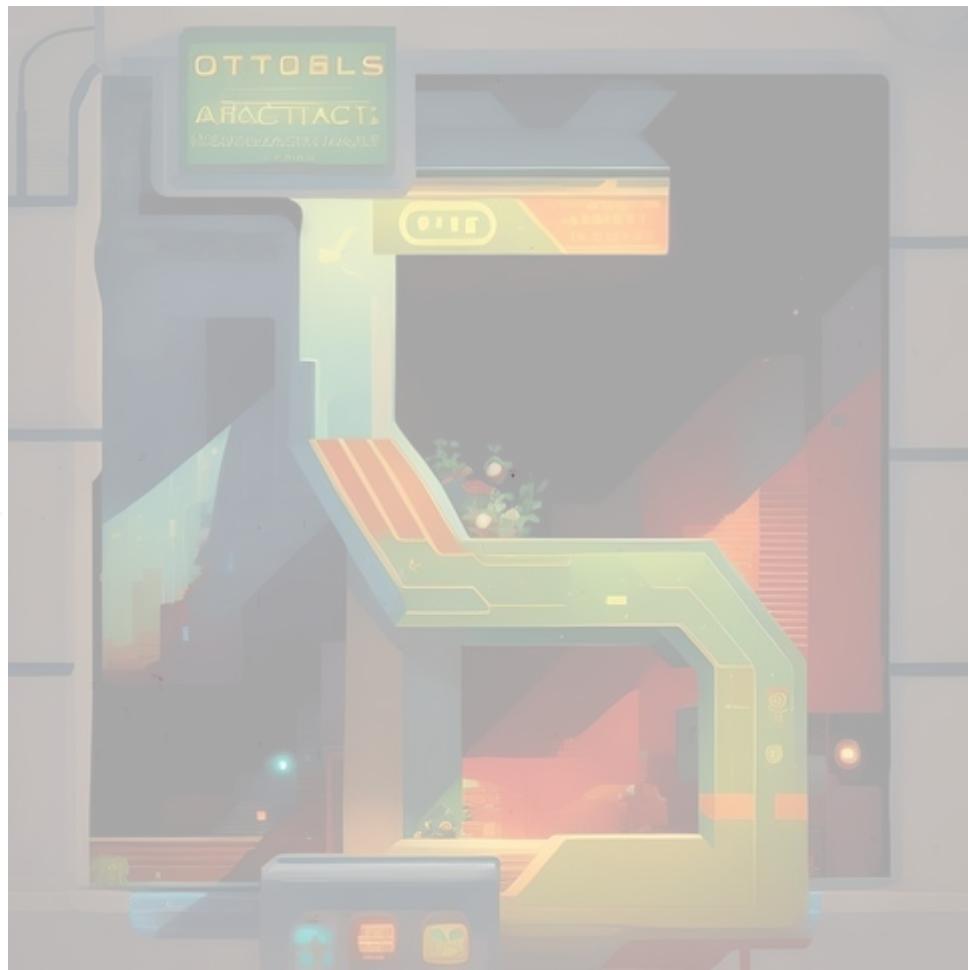
$$\therefore \text{I.C} = 16(1.4688) + 0.35(1.4688)^2$$

$$\begin{array}{r} 14400 \\ 14688 \\ \hline 512 \end{array}$$

$$\therefore \text{I.C} = 23.5008 + 0.7556$$

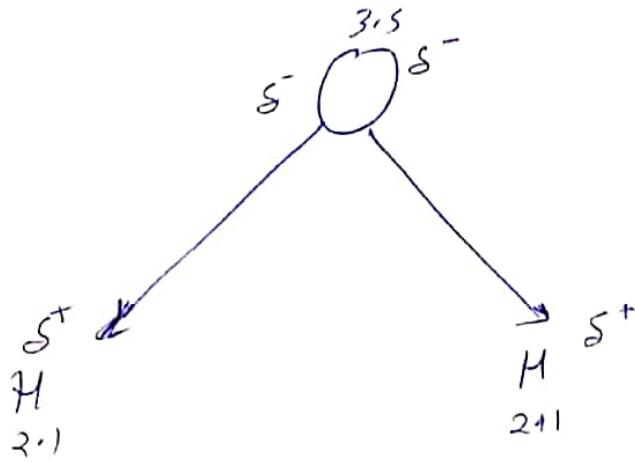
$$\boxed{= 24.25\%}$$

$$\text{Bond length} = D_{A,B} = 0.5 + 0.6 - 0.09(1.4688)$$
$$= 0.5 + 0.6 - 0.132 \cancel{1.4688}$$
$$= 0.968 \text{ \AA}^{\circ}$$



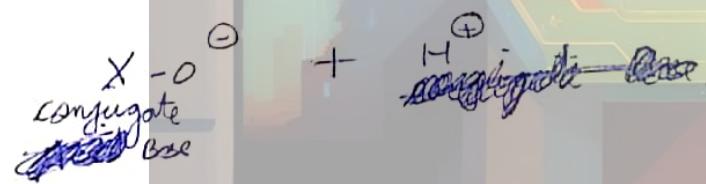
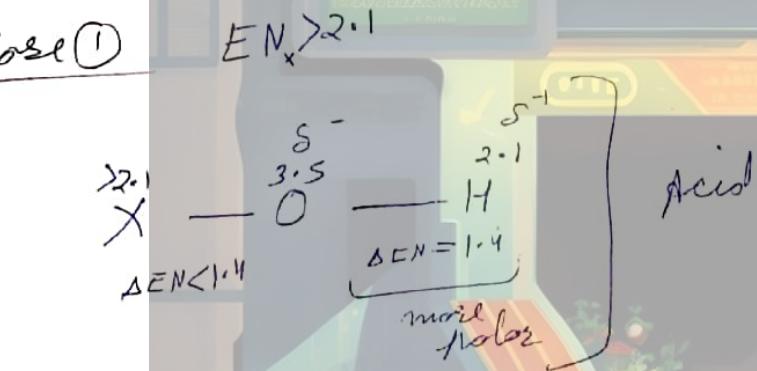
Nature of Hydroxide

→

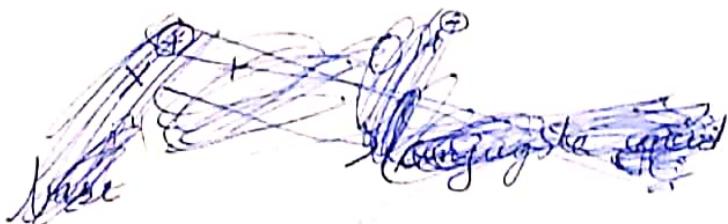
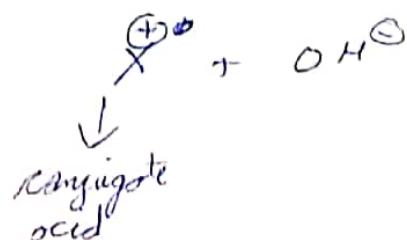
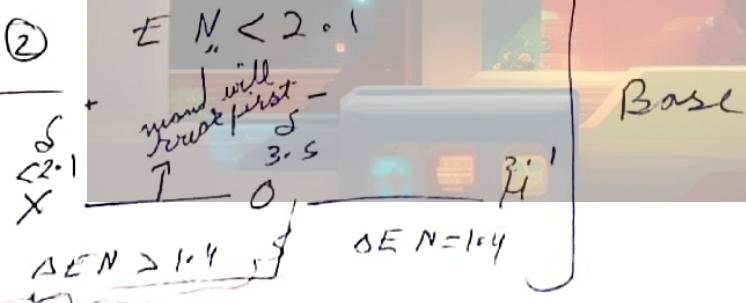


more polar bond will break in aqueous solution.

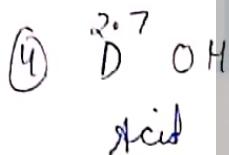
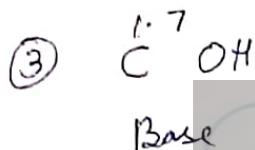
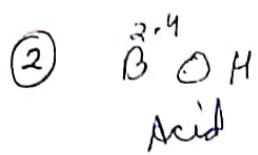
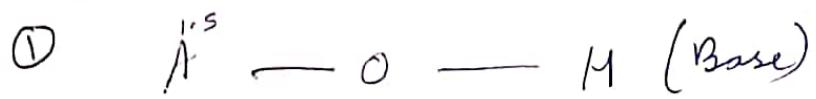
Case ①



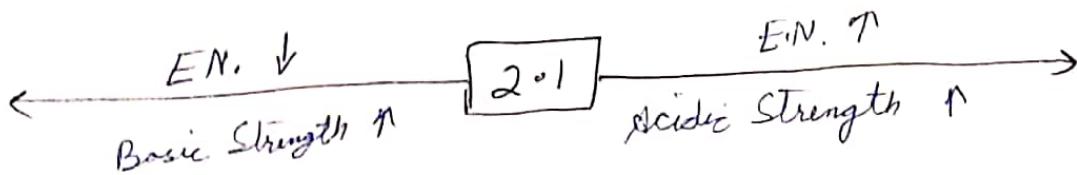
Case ②



Q find Acid & Base.

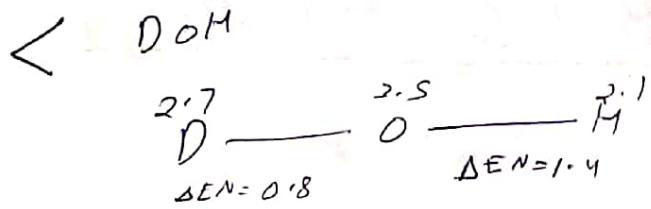
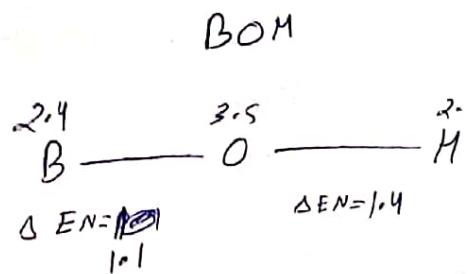


Acidic And Basic Strength of Hydroxides.



Ex.:

Acidic Strength



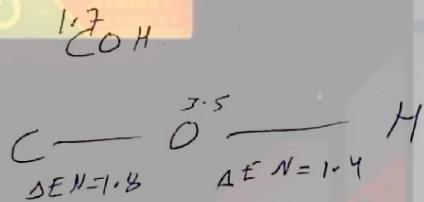
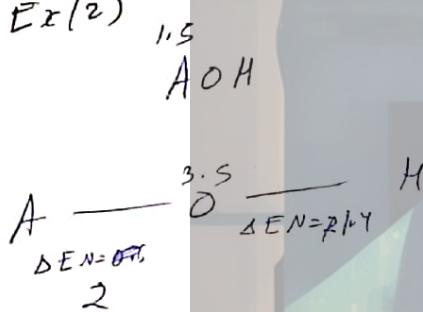
less ~~SEN~~, less strength in ~~bond~~
break easily

conjugate Base more ~~stable~~ stable

Thus, Acid strength ↑

Basic strength

Ex(2)



Break easily (because of high polarity)

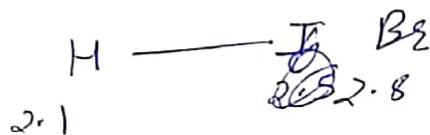
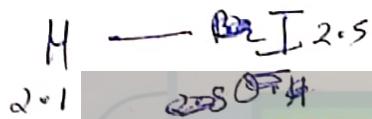
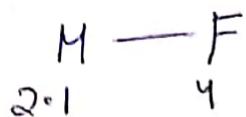
More Basic strength ↑

More stable Conjugate Acid.

Bond Strength/Bond Energy

Bond Strength $\propto \Delta E_{\text{EN}}$

Ex:-



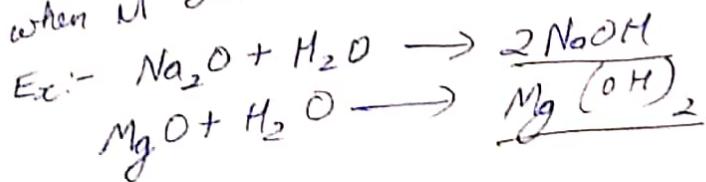
$\text{HF} > \text{HCl} > \text{HBr} > \text{HI}$ (Bond strength)
 $\Delta E_{\text{EN}} \uparrow$, bond strength \uparrow

$\text{HF} < \text{HCl} < \text{HBr} < \text{HI}$ (Acidic strength)
 $\Delta E_{\text{EN}} \downarrow$, acidic strength \downarrow

Types of Oxides

① Basic oxides - $(M^{\uparrow \text{ metal}})$ form basic oxides which reacts with acid.

→ when it dissolve in water (H_2O) form base.

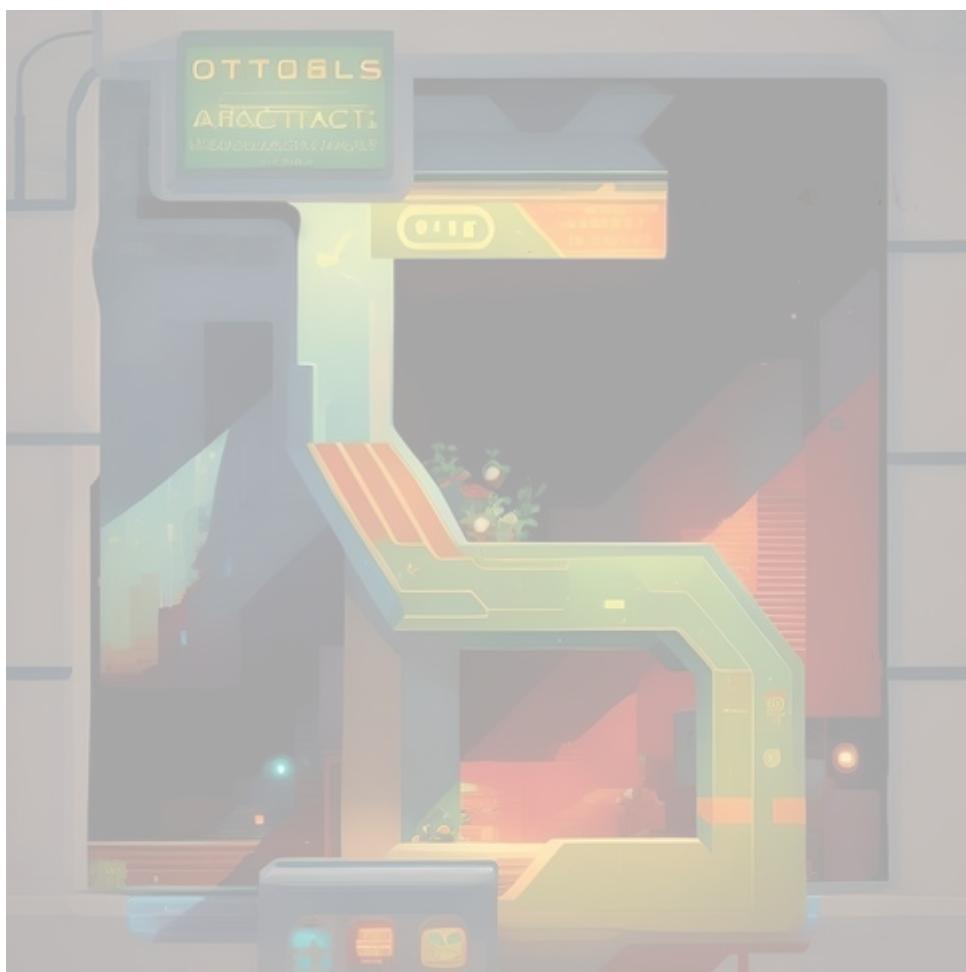


M.W.

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~~2024~~

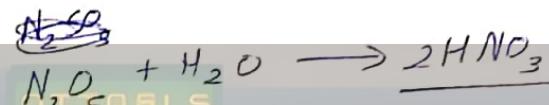
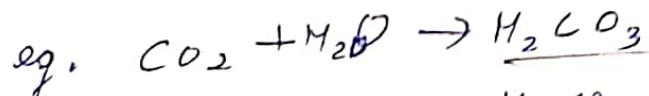
O-1, S-1 (complete)



② - Acidic Oxides ($\overset{\text{non metals}}{\underset{\text{N/M/A}}{\text{NM}}} \rightarrow \text{Acidic}$)

→ usually Non-metals forms Acidic oxides which reacts with base.

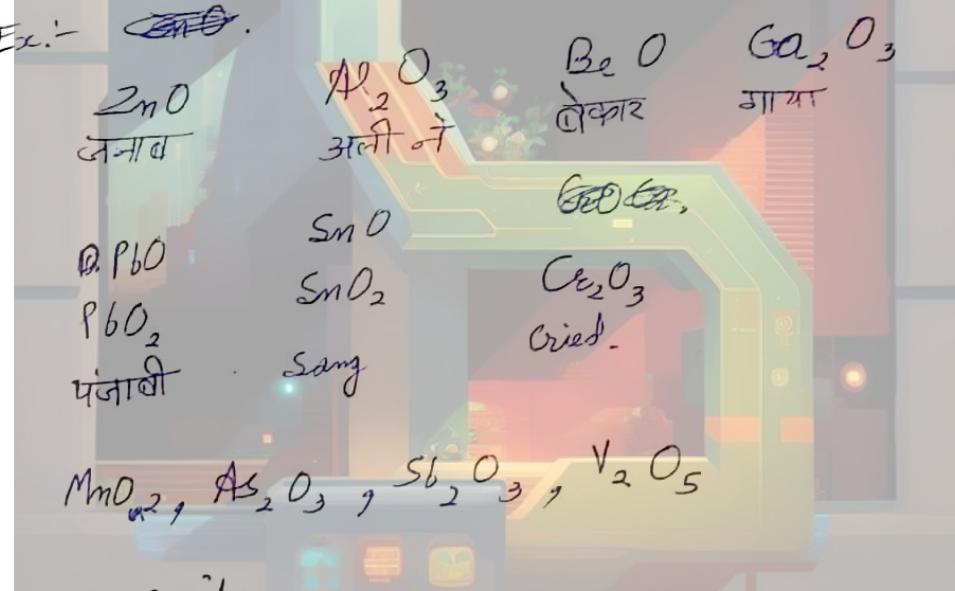
→ when it dissolves in water forms acids.



③ - Amphoteric Oxides -

→ They react with both acids and bases.

Ex. - ~~Al_2O_3~~ .



④ - Neutral Oxides

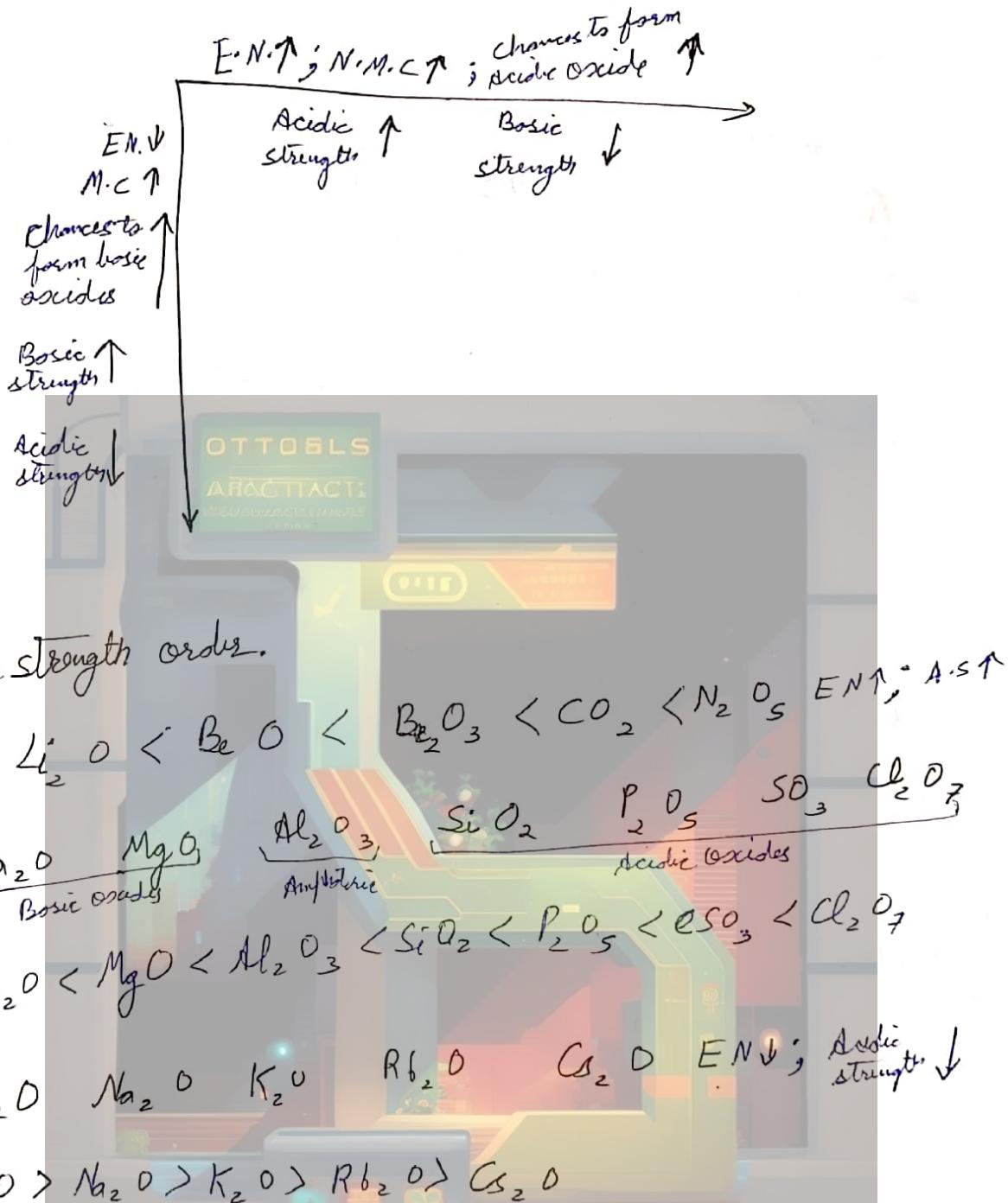
→ $\text{N}_2\text{O}, \text{NO}$

→ oxides which do not react with acid as well as base.

→ Ex. $\text{N}_2\text{O}, \text{NO}, \text{CO}, \text{H}_2\text{O}$ → (If treated as solvent, then considered neutral)

→ (If not used as a solvent, considered Amphoteric)

Periodicity of oxides



Oxidation Number

→ It is a hypothetical charge on element if every bond is considered to be ionic

Some general oxidation state -

① Alkali Metal ⇒ +1 {Li, N, K, Rb, Cs, Fr}

In Bonded state

② Alkali Earth Metals ⇒ +2 {Be, Mg, Ca, Sr, Ba}

② Hydrogen ⇒ Hydride with metal ⇒ -1 {LiH, NaH, KH...}

Hydrides with non metals / other elements ⇒ +1, {NH₃, PH₃, CH₄}

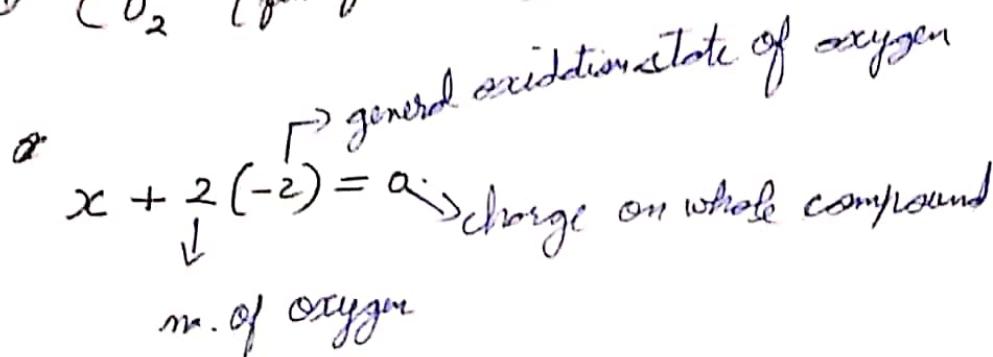
③ Oxygen general oxidation state ⇒ -2

④ Fluorine ⇒ -1 (Always)

Acidic Strength & oxidation state

H₂O₂ ⇒ -1
KO₂ ⇒ -½
KO₃ ⇒ -⅓
OF₂ ⇒ +2
O₂F₂ ⇒ +1

Q. ① CO_2 (find for carbon)



$$x - 4 = 0$$

$$\boxed{x = 4}$$

② SO_2

$$x + 2(-2) = 0$$

$$x - 4 = 0$$

$$\boxed{x = 4}$$

③ SO_4^{2-}

$$x + 4(-2) = -2$$

$$x - 8 = -2$$

$$\frac{x - 8}{x} = -2$$

$$\boxed{x = 6}$$

non metal

④ NH_3

$$x + 3(+1) = 0$$

$$x + 3 = 0$$

$$\boxed{x = -3}$$

⑥ H_2N_2

$$1 + 2x + 3(-2) = 0$$

$$2x + 1 - 6 = 0$$

$$\begin{cases} 2x = 6 - 1 \\ \boxed{x = 5} \end{cases} \checkmark$$

⑦ Na_2O_2

$$2x + 5(-2) = 0$$

$$2x - 10 = 0$$

$$2x = 10$$

$$x = \frac{10}{2}$$

$$\boxed{x = 5} \checkmark$$

⑧ FeO_4

$$x + 4(-2) = -3$$

$$x - 8 = -3$$

$$\begin{cases} x = 8 - 3 \\ \boxed{x = 5} \end{cases}$$

⑨ NH_4^+

$$x + 4(+1) = 1$$

$$x + 4 = 1$$

$$x = 1 - 4$$

$$\boxed{x = -3} \checkmark$$

(160)



$$x + 4(-2) = -1$$

$$x - 8 = -1$$

$$x = -1 + 8$$

$$\begin{cases} x = 8 - 1 \\ x = 7 \end{cases} \checkmark$$

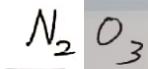
⑩



$$x + 4(+1) = 0$$

$$\begin{cases} x = -4 \\ \text{OTOBLS} \\ \text{ARCTACT} \end{cases} \checkmark$$

⑪



$$2x + 3(-2) = 0$$

$$2x - 6 = 0$$

$$2x = 6$$

$$\begin{cases} x = 3 \\ \text{OTOBLS} \end{cases} \checkmark$$

⑫

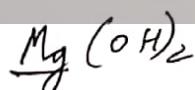


$$x + 3(+1) + 2(-2) = 0$$

$$x + 3 - 4 = 0$$

$$\begin{cases} x = 4 - 3 \\ x = 1 \end{cases} \checkmark$$

⑬



$$x + 2(-1) = 0$$

$$\begin{cases} x - 2 = 0 \\ x = 2 \end{cases} \checkmark$$

⑯

(14) Na

$$\boxed{x=0} \checkmark \text{ (Elemental State) (Non-Bonded)}$$

(15) O_2

$$2x \neq 0$$

$$2(-2) = -4$$

$$\boxed{\begin{array}{l} 2x=0 \\ x=0 \end{array}} \checkmark$$

(16) H_2O_2

$$2x + 2(+1) = 0$$

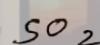
$$2x + 2 = 0$$

$$x = \frac{-2}{2}$$

$$\boxed{x=-1} \checkmark$$

Q Compare Acidic Strength

Ex ①



$$x + 2(-2) = 0$$

$$x = +4$$



$$x = 6$$

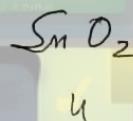
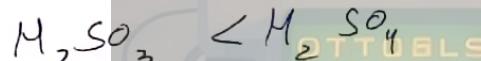
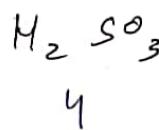
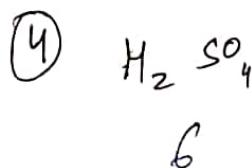
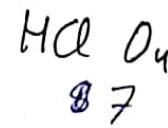
$$\boxed{\text{SO}_2 < \text{SO}_3} \checkmark$$

(2)

N_2O	NO	N_2O_3	NO_2	N_2O_5
1	2	3	4	5

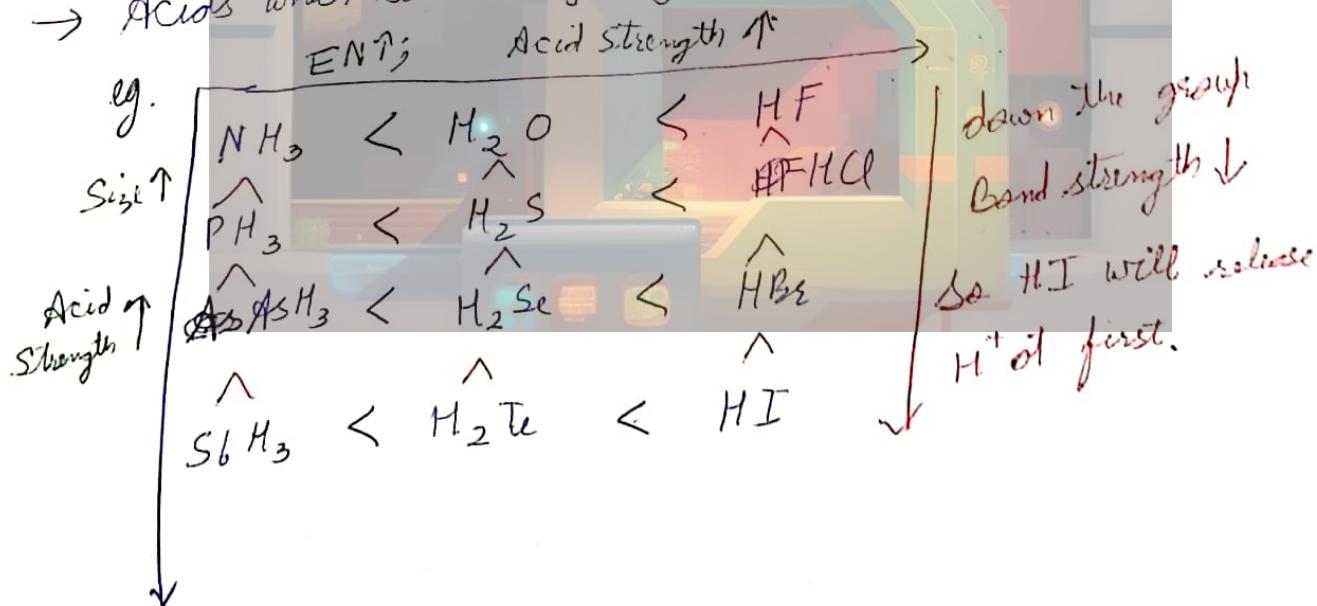
$$\boxed{\text{N}_2\text{O} < \text{NO} < \text{N}_2\text{O}_3 < \text{NO}_2 < \text{N}_2\text{O}_5} \checkmark$$

(152)



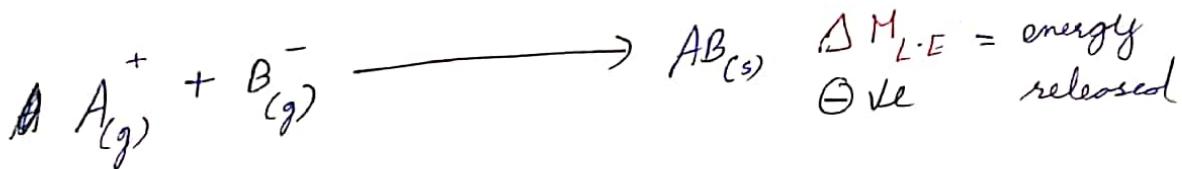
Hydro Acids

→ Acids which contains hydrogen with other non metal

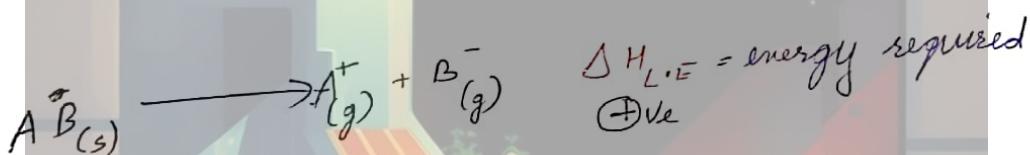


* Lattice Energy ($\Delta H_{L.E.}$)

→ It is the amount of ~~excess~~ energy released during the formation of 1 mole lattice/solid from constituent gaseous ions.



→ It is the amount of energy required to break one mole solid/lattice ^{the into their constituent} gaseous ions.



factors of L.E. :-

$$L.E \propto q_1 q_2 \quad \left\{ \begin{array}{l} q_1 \rightarrow \text{charge on cation} \\ q_2 \rightarrow \text{charge on anion} \end{array} \right\}$$

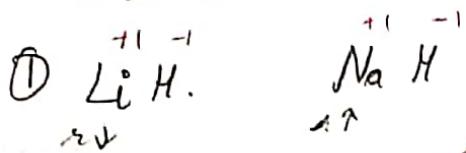
$$L.E \propto \frac{1}{r} \quad \left\{ r \rightarrow \text{internuclear distance} \right\}$$

$$r \approx r_A + r_C \quad \begin{matrix} \downarrow \\ \text{radius of} \\ \text{anion} \end{matrix} \quad \rightarrow \text{product of cation}$$

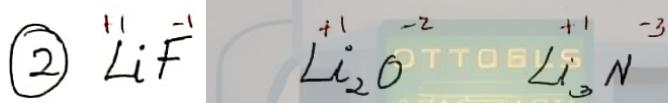
if $|q_1 q_2|$ or amount of charge ↑; $L.E. \uparrow$
 if $r \uparrow$; $L.E. \downarrow$

(9, 9₂) P; r↑; L.E.↑ { Phle charge & ka
decreas for radius ka }

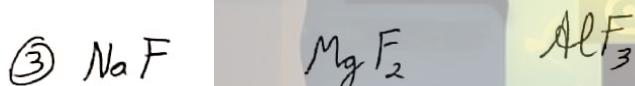
Q1. L.E. order?



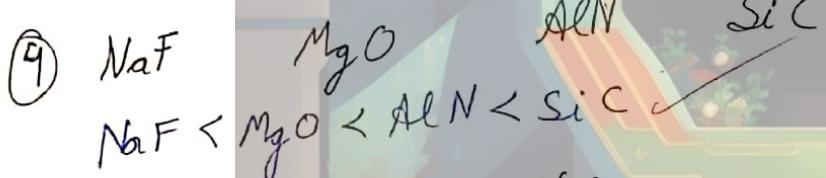
$$\text{LiH} > \text{NaH} \checkmark$$



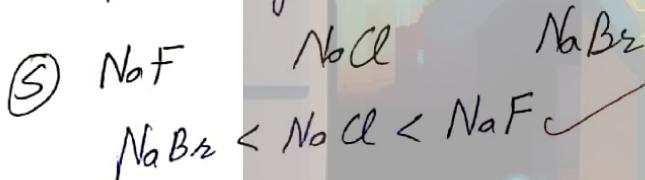
$$\text{LiF} < \text{Li}_2\text{O} < \text{Li}_3\text{N} \checkmark$$



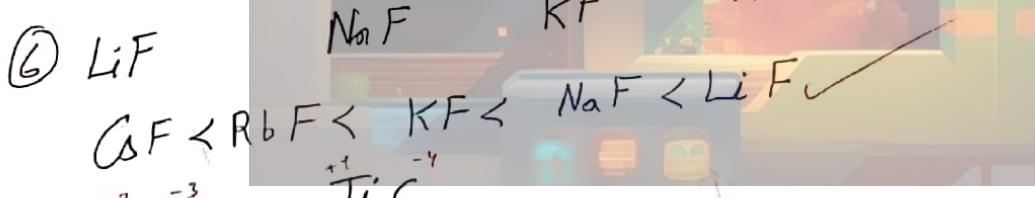
$$\text{NaF} < \text{MgF}_2 < \text{AlF}_3$$



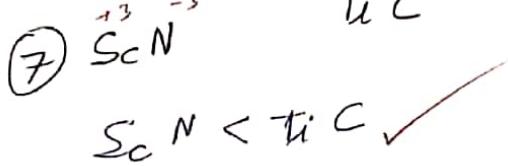
$$\text{NaF} < \text{MgO} < \text{AlN} < \text{SiC}$$



$$\text{NaBr} < \text{NaCl} < \text{NaF}$$



$$\text{CsF} < \text{RbF} < \text{KF} < \text{NaF} < \text{LiF}$$



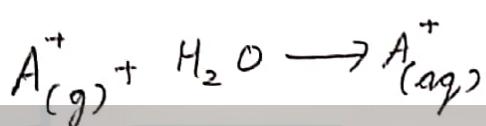
$$\text{ScN} < \text{TiC}$$

$\text{N}^{3-} \rightarrow \text{Nitride}$

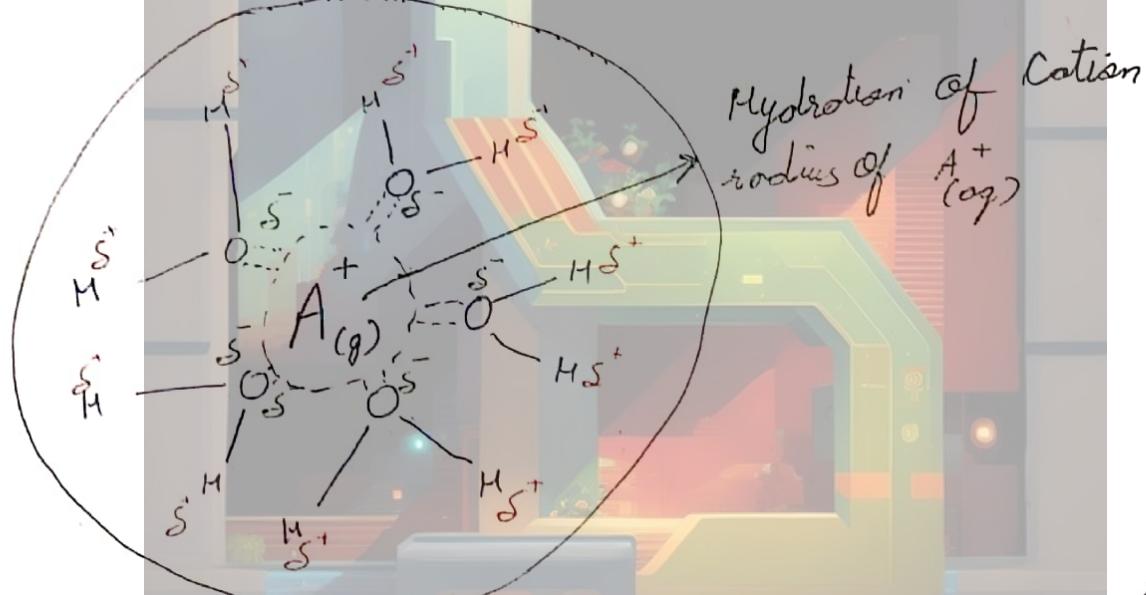
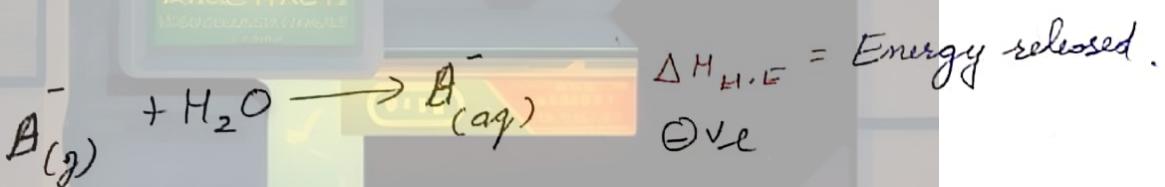
$\text{C}^{4-} \rightarrow \text{Carbide}$

Hydration Energy

→ It is the amount of energy released when 1 mole of gaseous ions are interested with water (dissolve)



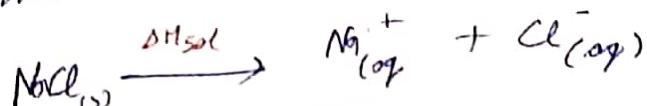
$\Delta H_{H.E}$ = Energy released
Over



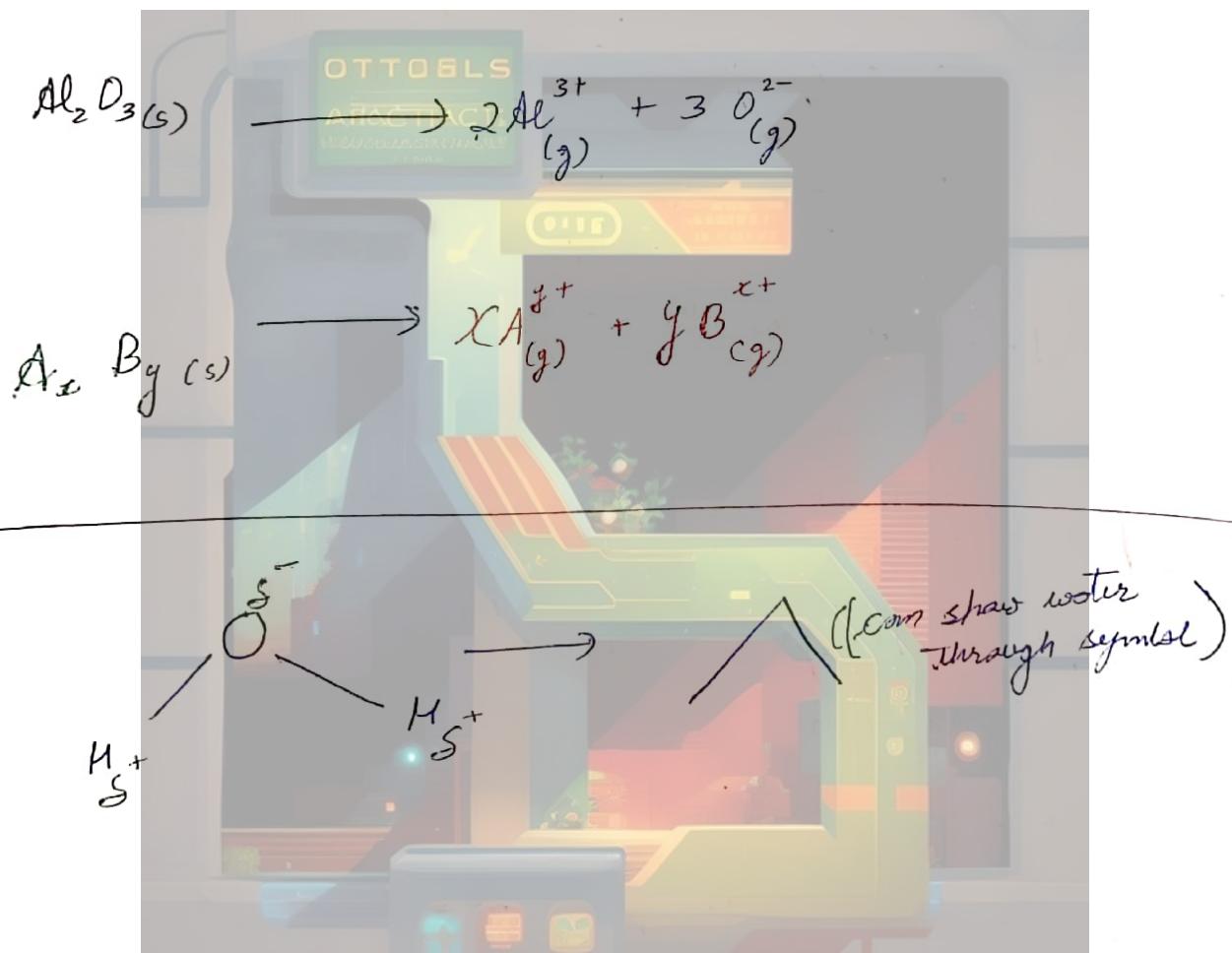
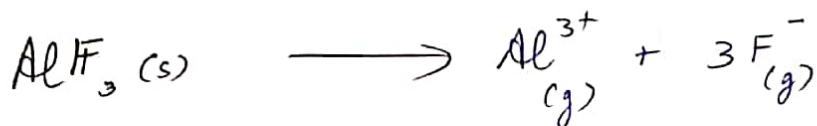
More water molecule interact ↑; \therefore hydration energy ↑.

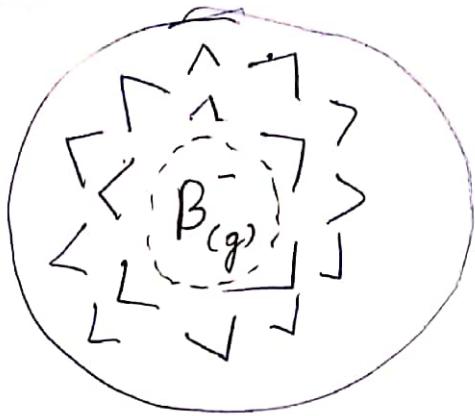
radius of aqueous ↑

Enthalpy of solution - (ΔH_{sol}) - amount of energy when 1 mole of solid is converted to constituent aqueous ions.



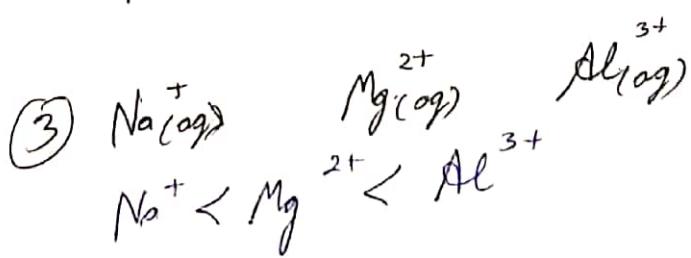
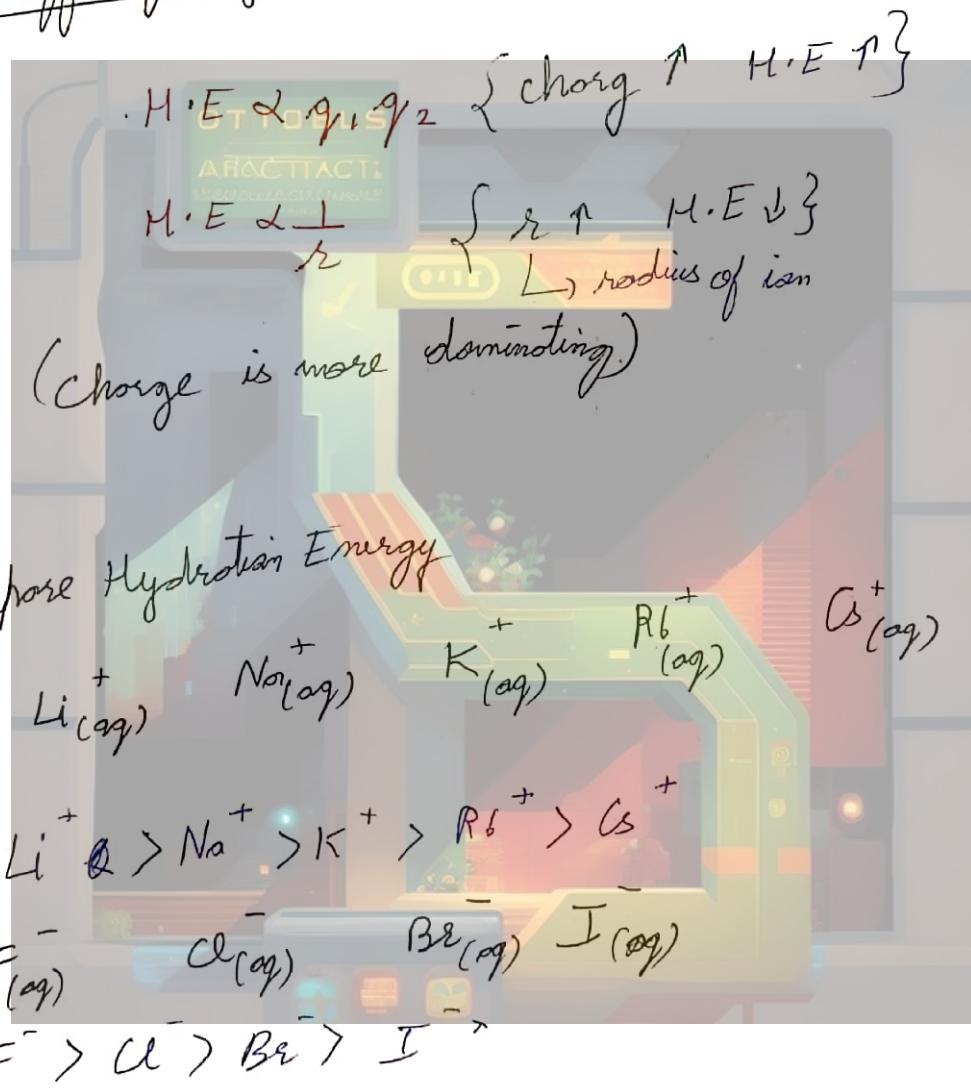
$$\Delta H_{sol} = \Delta H_{L.E} + \Delta H_{H.E}$$

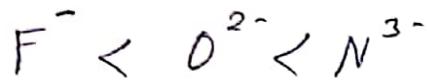
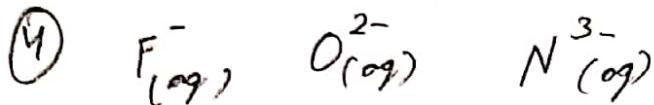




Hydration of anion

Factors affecting Hydration Energy



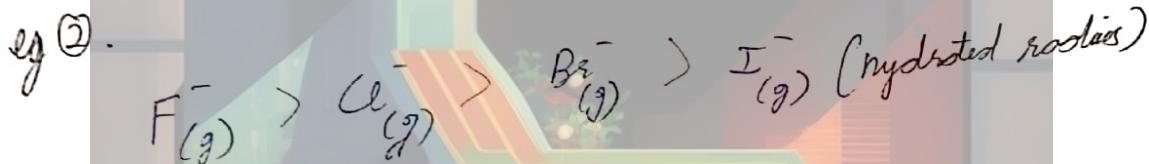
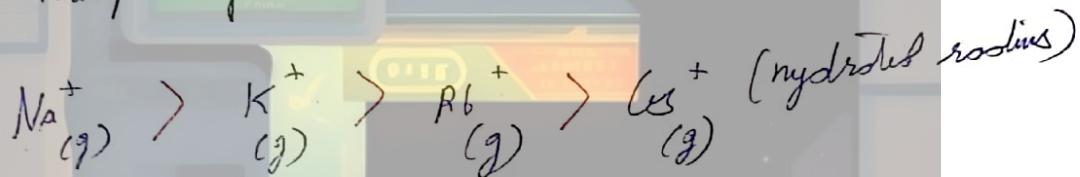


Application of hydration energy :-

① Hydrated radius -

$H \cdot E. \uparrow$; Number of water molecules \uparrow Hydrated Radius \uparrow

e.g. Compare hydrated radius



Hydrated Radius $\propto H \cdot E.$

② Ionic Mobility -
 \rightarrow Ionic mobility means movement of ions.

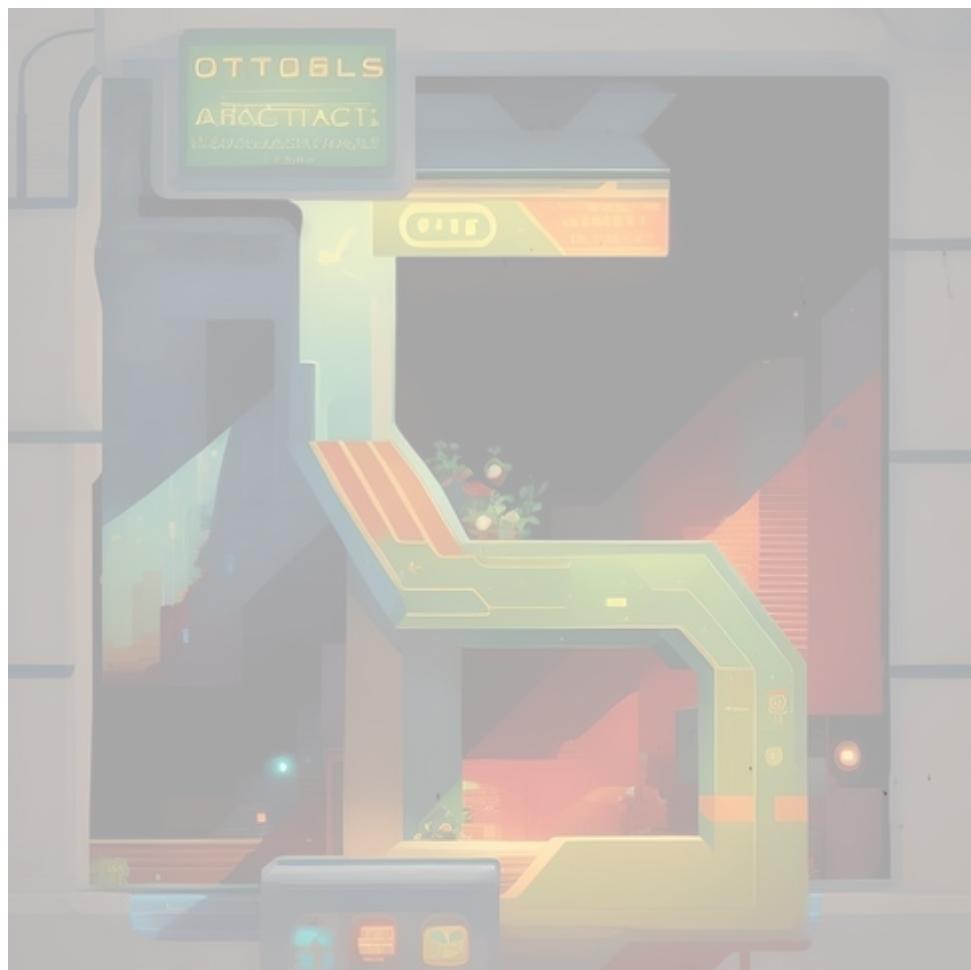
$H \cdot E. \uparrow$; more no. of H_2O interact \uparrow

Hydrated radius \uparrow ; Movement of ions \downarrow

~~H.E.~~ $\boxed{\text{Ionic Mobility} \propto \frac{1}{H \cdot E.}}$

H.W. 11-05-2024

O-Z, Mains & advance full



③ Electric Conduction

Ionic mobility & Electrical conductivity.

H.E.P Hydrated Radius ↑ Ionic mobility ↓ Electrical conductivity ↓

Conduction - Flow of Charge per unit time.

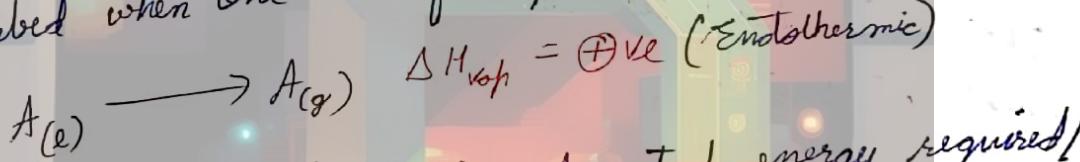
* Bond En-Haber Cycle

Imp Terms -

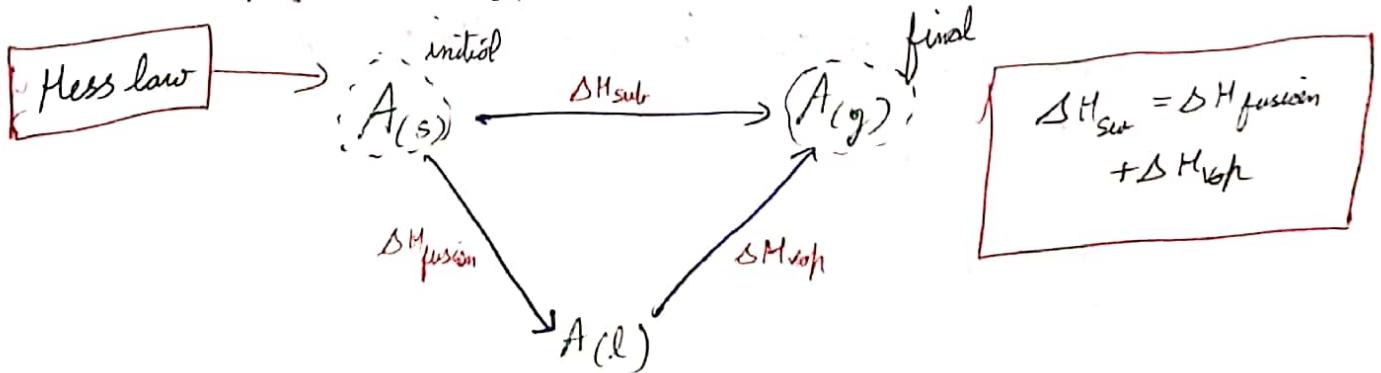
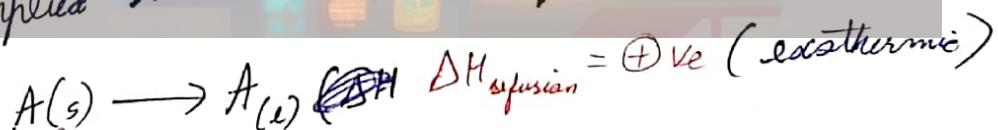
① Enthalpy of Sublimation (ΔH_{sub}) - Amount of energy required/ absorbed when 1 mole of solid is converted to gas.



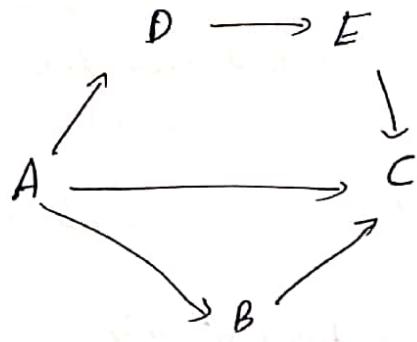
② Enthalpy of Vapourisation (ΔH_{vap}) - Amount of energy supplied/ absorbed when one mole of liquid is converted into gas.



③ Enthalpy of Fusion (ΔH_{fusion}) - Amount of energy required/ supplied to convert 1 mole of solid into liquid.



Hess Law-



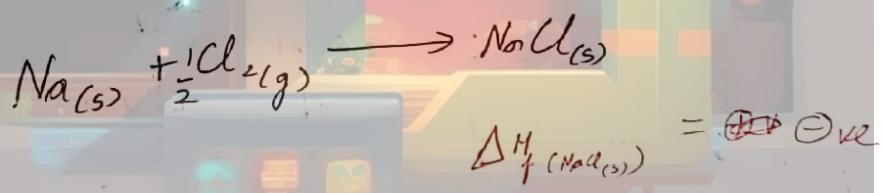
$$\Delta H_{AC} = \Delta H_{ADEC} = \Delta H_{ABC}$$

$$\Delta H_{AC} = \Delta H_{AO} + \Delta H_{DE} + \Delta H_{CE}$$

$$\Delta H_{AC} = \Delta H_{AB} + \Delta H_{BC}$$

* Enthalpy only calculated from final & initial points & path not considered.

④ Enthalpy of formation (ΔH_f) - It is the change in enthalpy when 1 mole of ionic solid is formed by their constituent elements in their reference state/standard state.



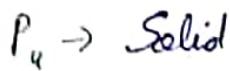
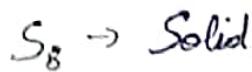
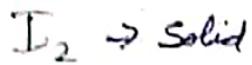
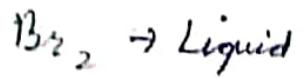
Standard states -

Metal \Rightarrow Solid State (Exception - Hg (Mercury) \Rightarrow Liquid)

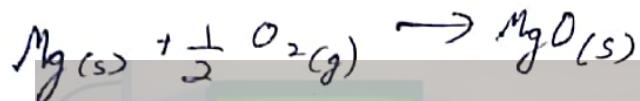
$O_2 \Rightarrow$
 $Cl_2 \Rightarrow$
 $N_2 \Rightarrow$
 $H_2 \Rightarrow$

Gas

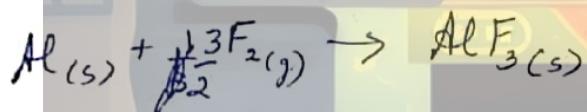
(162)



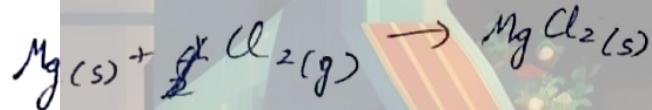
EgQ. ① $\Delta H_f (\text{MgO}_{(s)})$



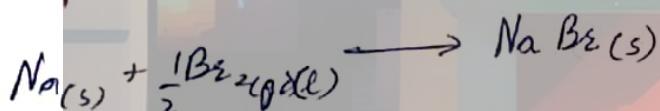
② $\Delta H_f (\text{AlF}_{3(s)})$



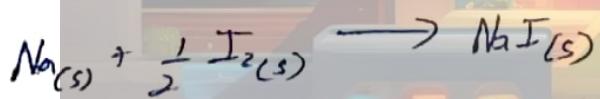
③ $\Delta H_f (\text{MgCl}_{2(s)})$



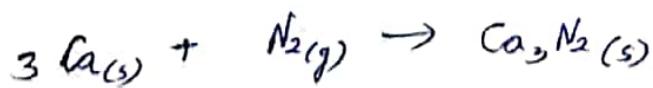
④ $\Delta H_f (\text{NaBr}_{(s)})$



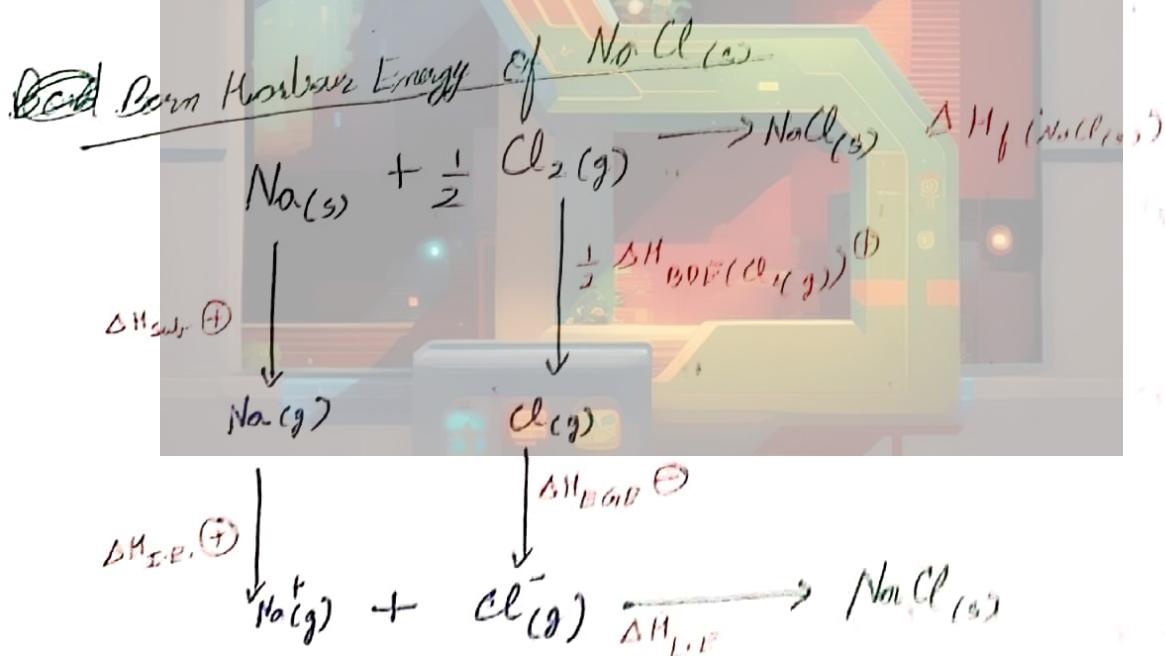
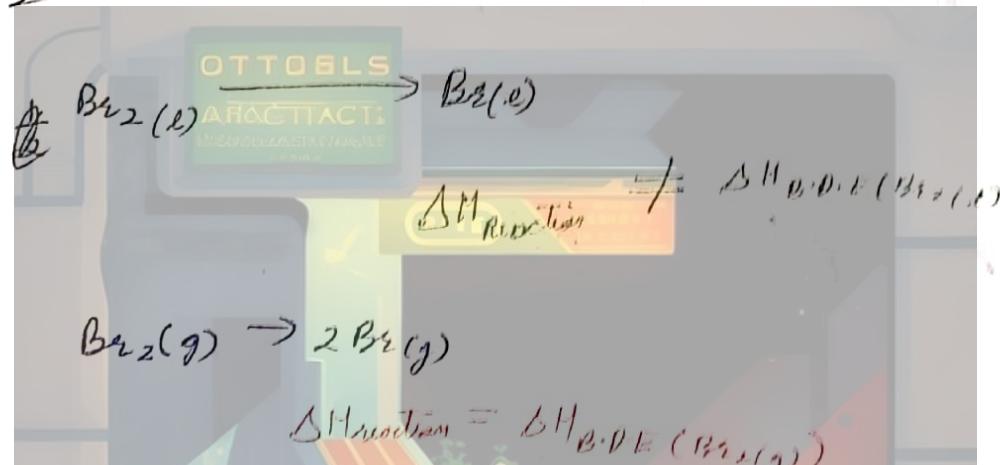
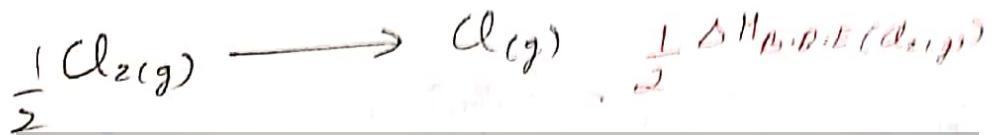
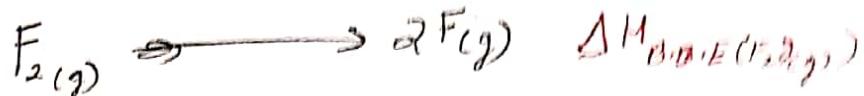
⑤ $\Delta H_f (\text{NaI}_{(s)})$



⑥ $\Delta H_f (\text{Ca}_3\text{N}_{2(s)})$



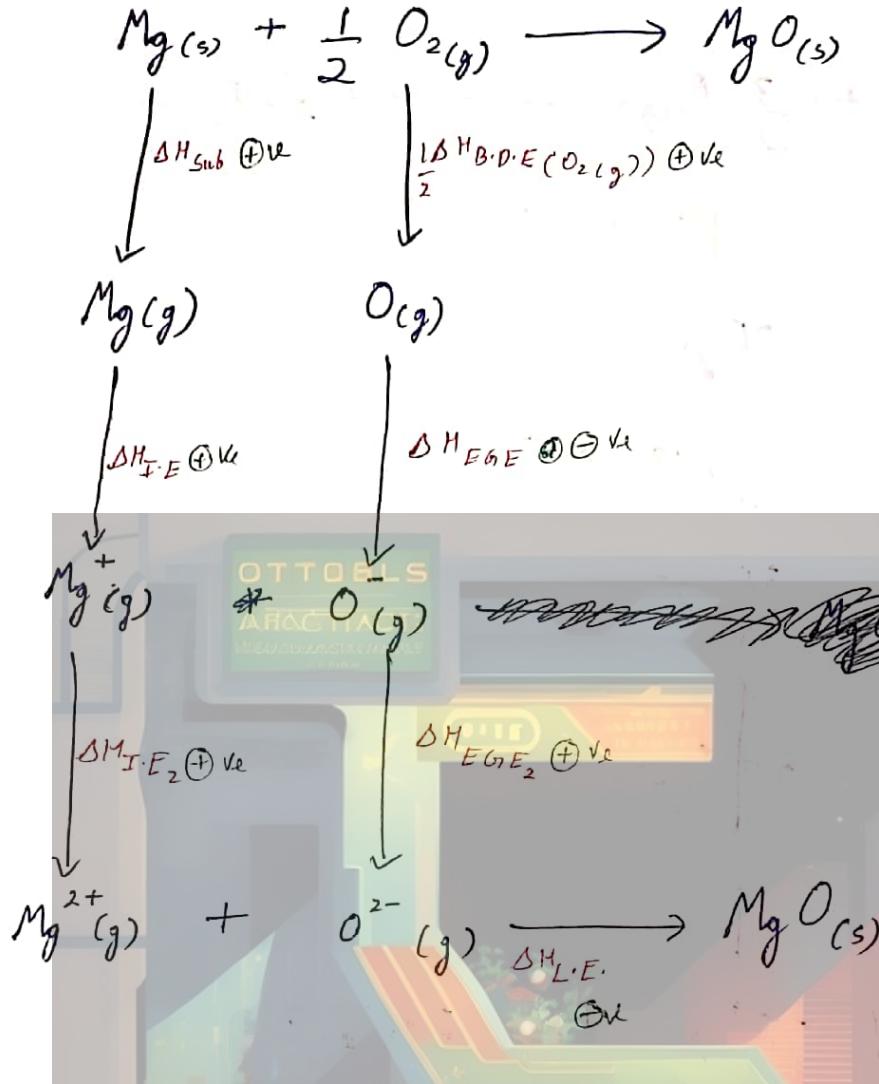
⑤ Bond Dissociation energy (B.D.E) - Amount of energy supplied to break one mole bond of gaseous molecules into their constituent gaseous atoms.



$$\Delta H_f(NaCl_{(s)}) = \Delta H_{\text{sub}(Na_{(s)})} + \Delta H_{I.P.}(Na_{(g)}) - \frac{1}{2} \Delta H_{B.D.E}(Cl_{2(g)}) + \Delta H_{B.G.B}(Cl_{(g)})$$

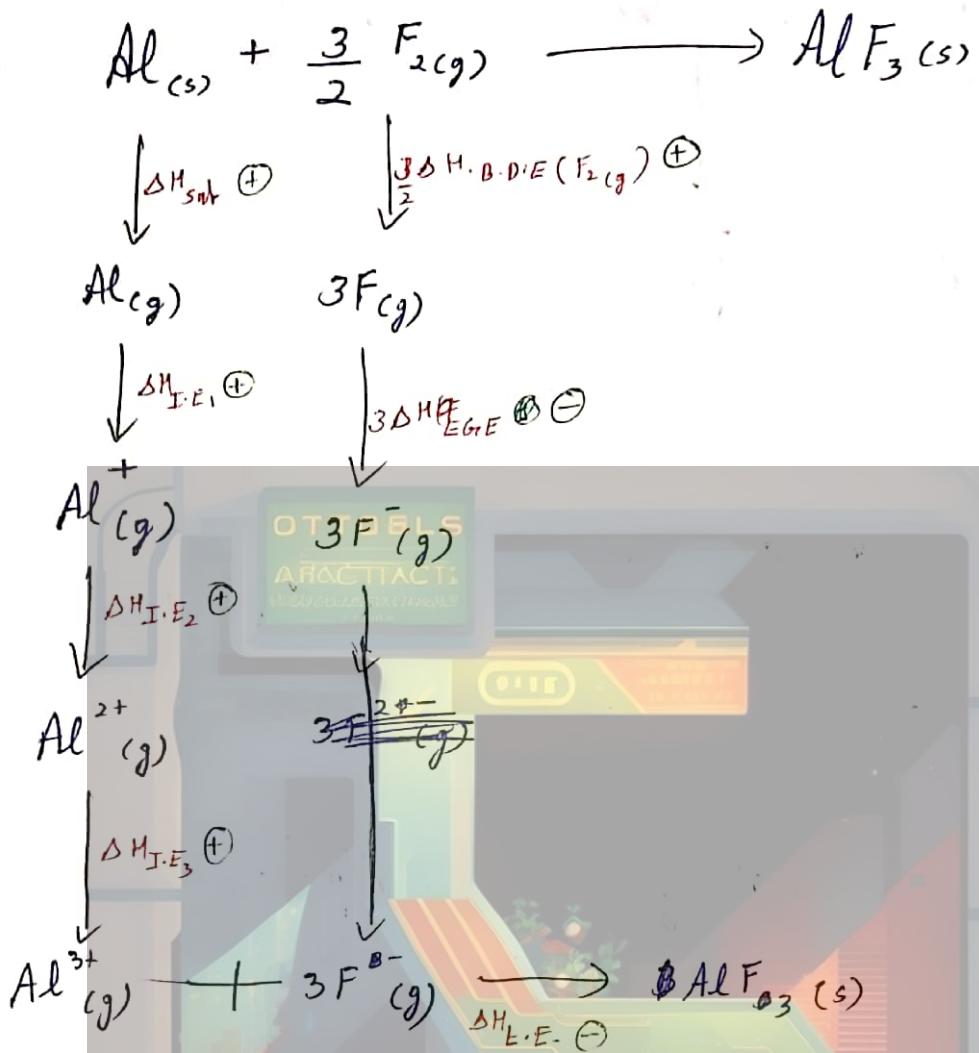
$$+ \Delta H_{L.F.}(NaCl_{(s)})$$

(2) write Born-Haber cycle of $MgO_{(s)}$



$$\begin{aligned}
 \Delta H_f(MgO_{(s)}) = & \Delta H_{\text{Sub}}(Mg_{(s)}) + \Delta M_{I.E.}(Mg_{(g)}) + \Delta H_{I.E.}(Mg^+_{(g)}) \\
 & + \frac{1}{2} \Delta H_{\text{B.O.E}}(O_{2(g)}) + \Delta H_{E.G.E}(O_{(g)}) + \Delta H_{E.G.E_2}(O^+_{(g)}) \\
 & + \Delta H_{L.E.}
 \end{aligned}$$

Q Born Haber cycle for $\text{AlF}_3(s)$

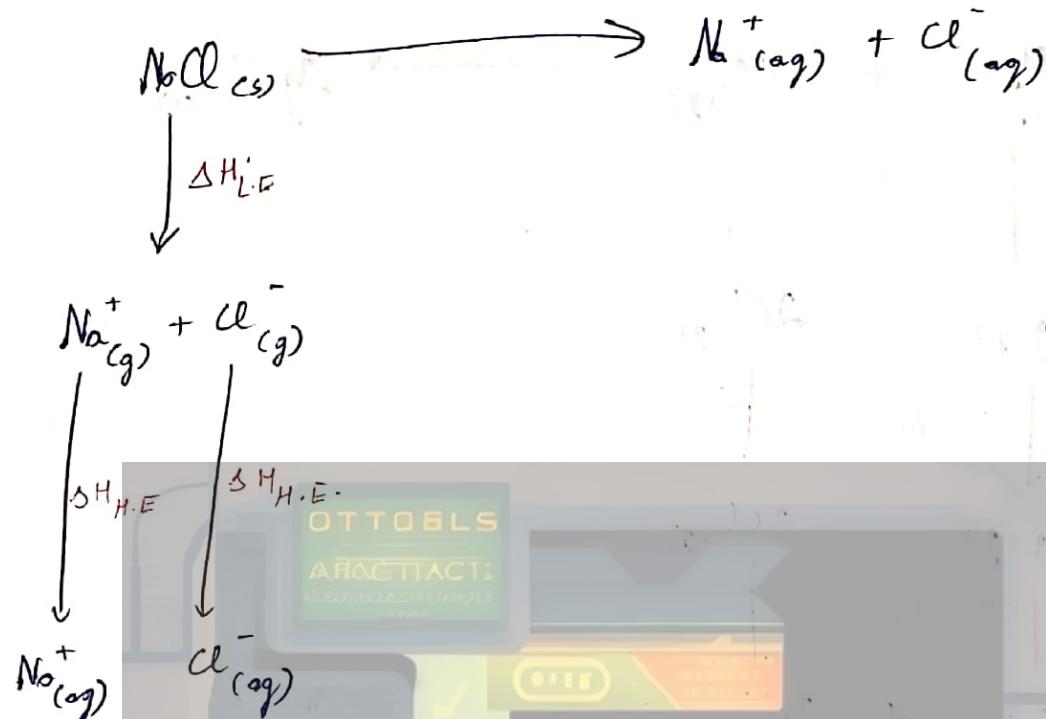


Q ④ $\text{MgCl}_2(s)$ H.W. 13-0 S-2024

⑤ $\text{AlN}(s)$ Rose-Till ST
S-2

⑥ $\text{Ca}_3\text{N}_2(s)$

Solubility of Ionic compound in H_2O



$$|\Delta H_{L.E.}| > |\Delta H_{H.E.}|$$

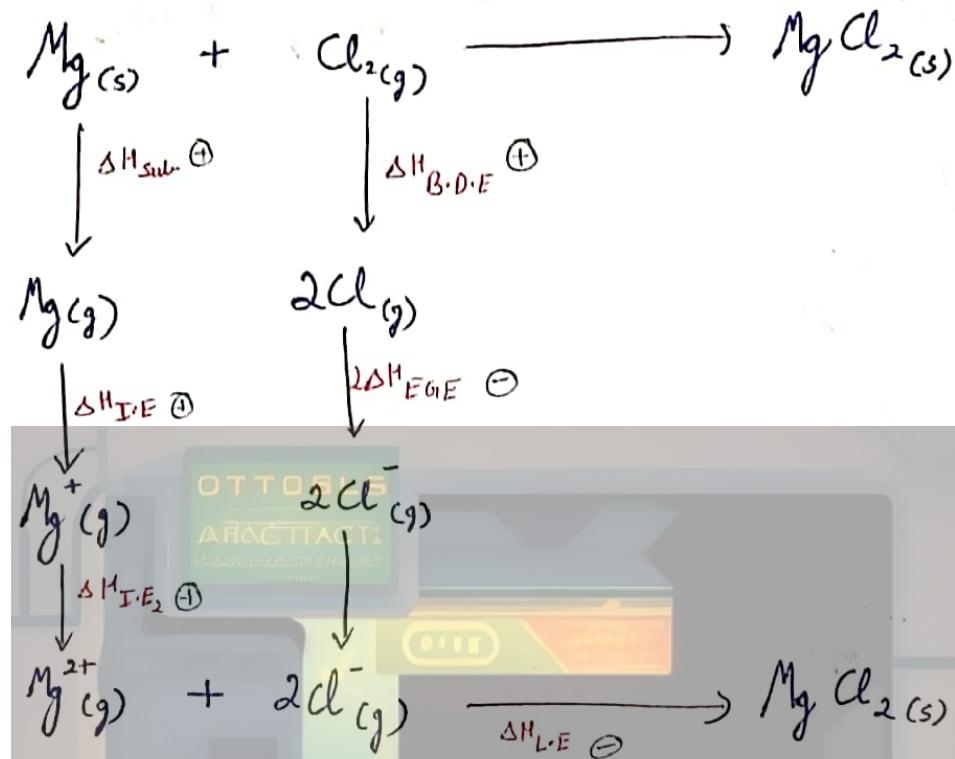
Not Soluble

$$|\Delta H_{L.E.}| < |\Delta H_{H.E.}|$$

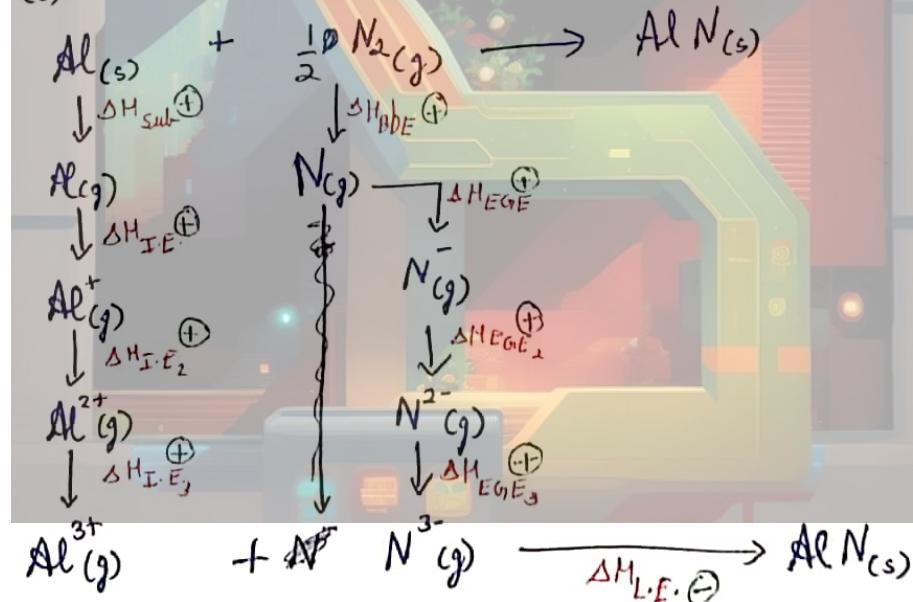
Soluble in H_2O

H.W.

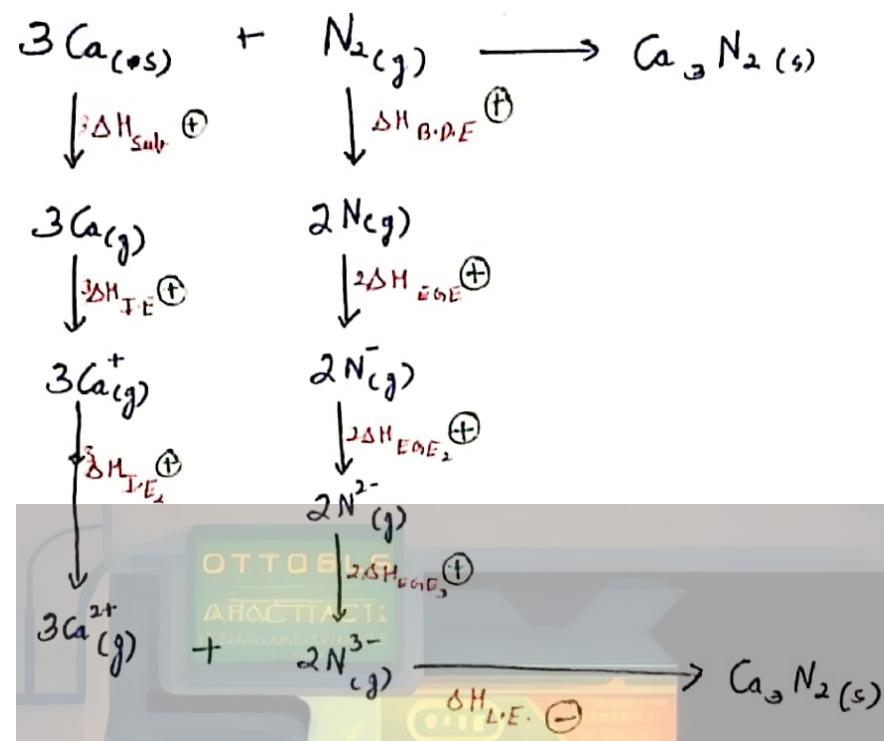
① $MgCl_2$



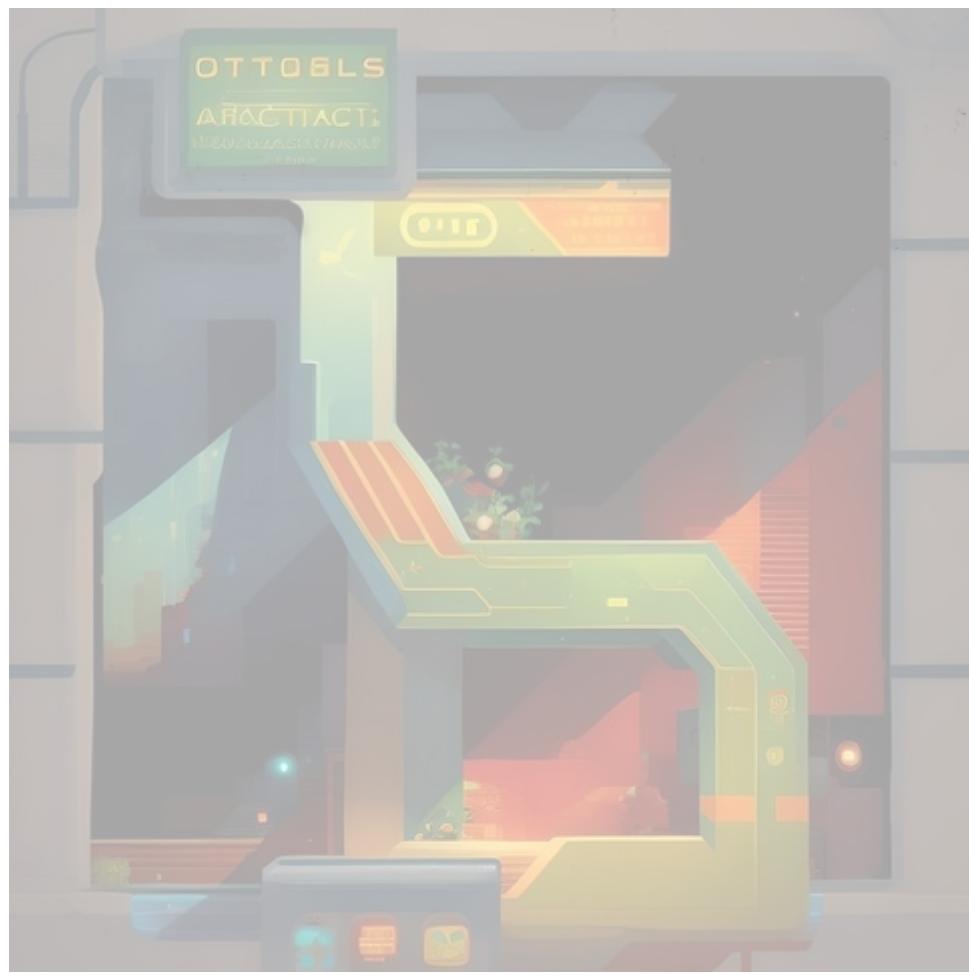
② $AlN_{(s)}$



③ Ca_3N_2



(169)

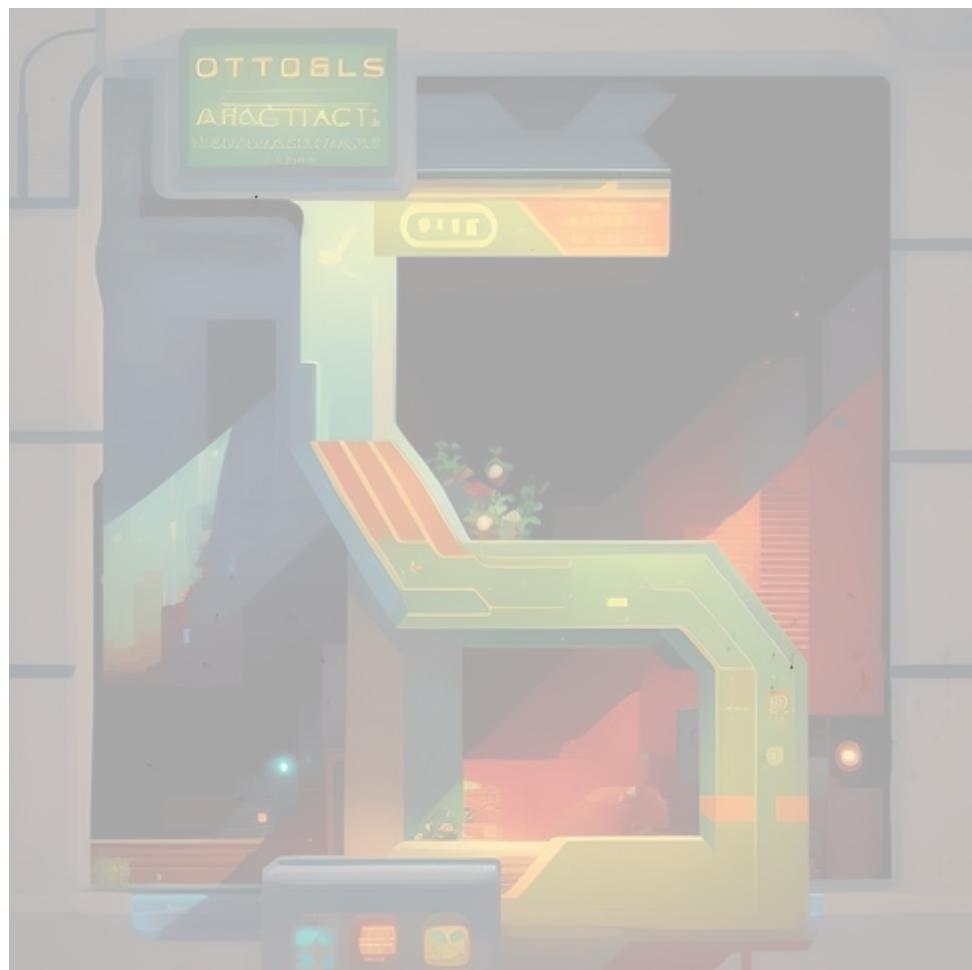




(17)



(172)



Chemical Bonding

Chemical bond - It is the force of attraction between two atoms which binds them together.

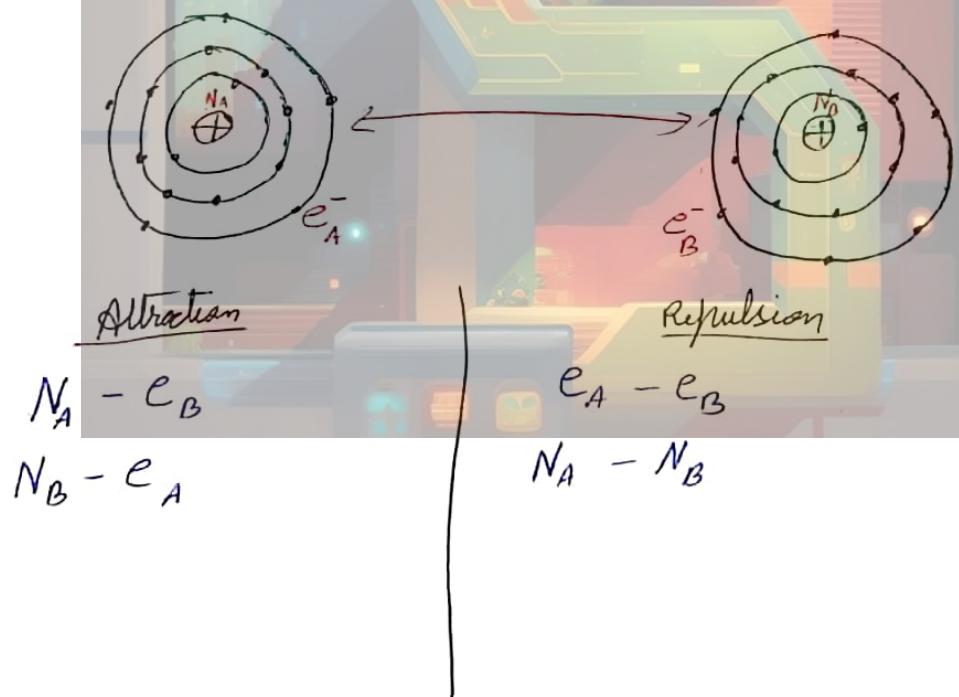
Why a chemical bond is formed?

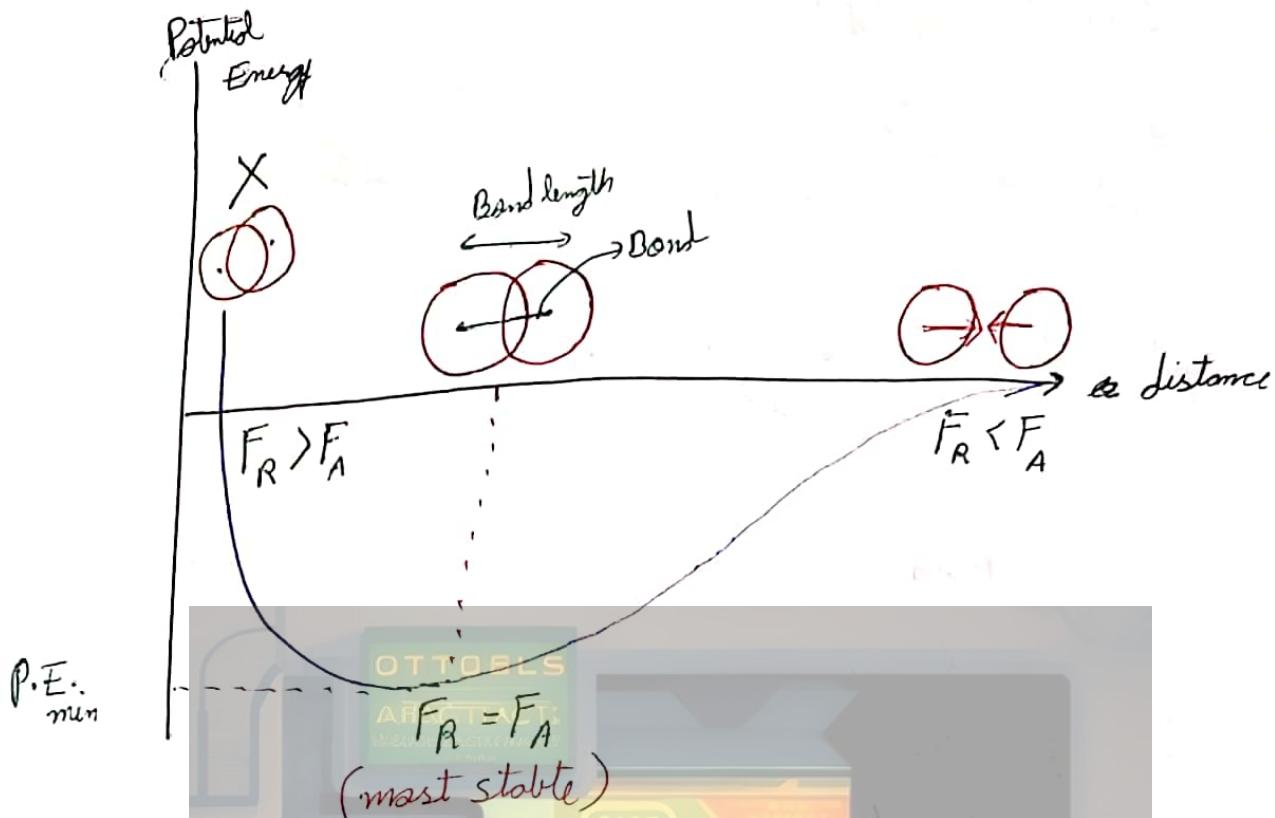
→ To gain stability by decreasing their potential energy or to acquire inert gas configuration.

→ Atoms are more reactive than molecules because they have higher energy than molecules.

$\text{Cl} \rightarrow$ atom (more reactive)

$\text{Cl}_2 \rightarrow$ molecule (less reactive)





F_A = Force of Attraction

F_R = Force of Repulsion

$$F_R = F_A$$

(most stable)

- When atoms approaches to each other, attraction force dominates over repulsion.
- Bond formation takes place when attraction force is equal to repulsion force because at this point, potential energy is minimum & stability is max.

Bond Energy

- It is ~~the~~ the amount of energy required to break bond.
- or
- It is the amount of energy released during bond formation.

Bond Length (r_0/d)

→ It is the shortest internuclear distance where bond formation takes place between two atoms.

Types of Chemical Bonds.

Strong Bond

($E_{\text{Released}} > 42 \text{ kJ/mol}$)

- ① Ionic / Electrovalent Bond
- ② Covalent Bond
- ③ Co-ordinate Bond / Active Bond
- ④ Metallic Bond.

Weak Bond

($E_{\text{Released}} < 42 \text{ kJ/mol}$)

Hydrogen Bond

(~~E_{Released}~~ & $E_{\text{Rel}} < 42$)

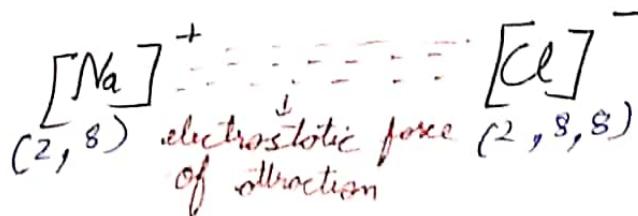
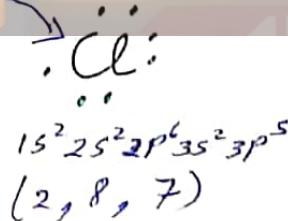
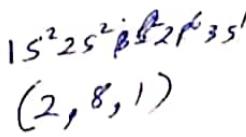
Weak forces

($E_{\text{Rel}} < 8 \text{ kJ/mol}$)

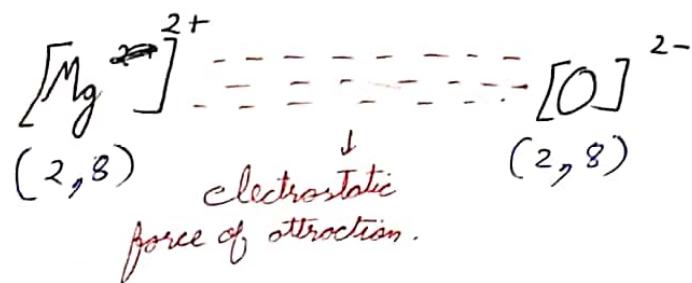
1) Ionic / Electrovalent Bond :-

→ It is the electrostatic force of attraction between two oppositely charged ions which are created by complete transfer of electron.

→ Eg:-



Eg 2.



XX → Ionic bond is non-directional in Nature.

Electrovalency → It is defined only for ionic bonds.

→ It is the number of electron transfer during formation of ionic bond by an atom.

Eg 1. NaCl E.v. (Na^+) = 1
E.v. (Cl^-) = 1

Ex 2. MgO E.v. (O^{2-}) = 2
E.v. (Mg^{2+}) = 2

3. AlF_3 E.v. (Al^{3+}) = 3
E.v. (F^-) = 1



Condition for the formation of ionic bond →

$$\rightarrow \boxed{\Delta E \cdot N \geq 2 \cdot 1}$$

- Low ionisation energy of electropositive element.
- High Electron Affinity of electronegative element.
- High Lattice energy.

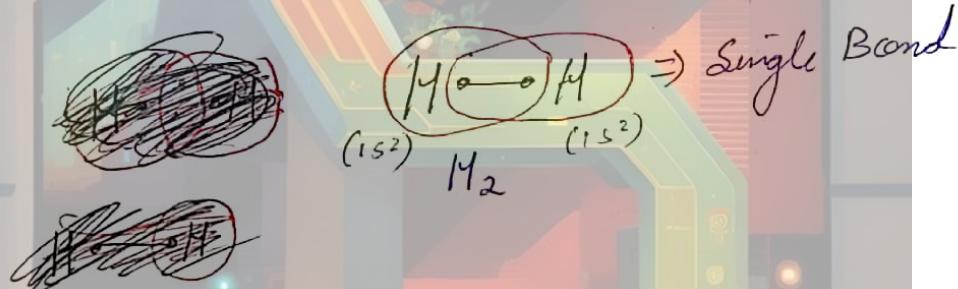
2. Covalent Bond:

→ This type of Bond is formed by equal sharing of electrons between two bonded atoms.

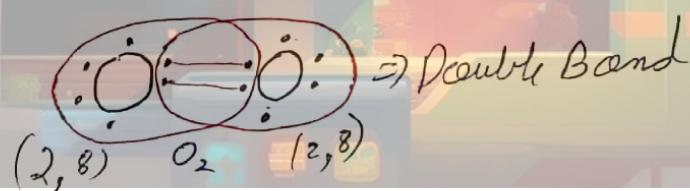
→ ~~Covalent Bonds are also~~

→ Covalent Bonds are directional in nature.

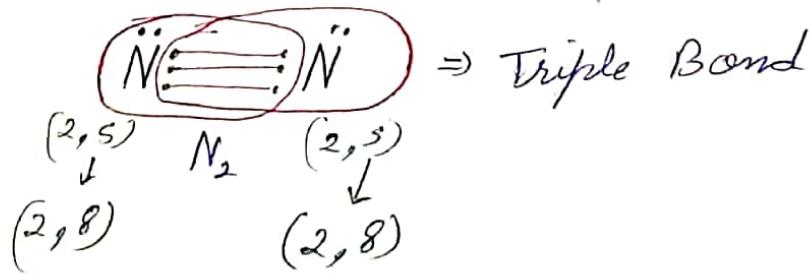
→ Eg 1.



Eg 2.



Eg 3.



Electron Pair

→ Pair of e^- is known as electron pair

Electron Pairs

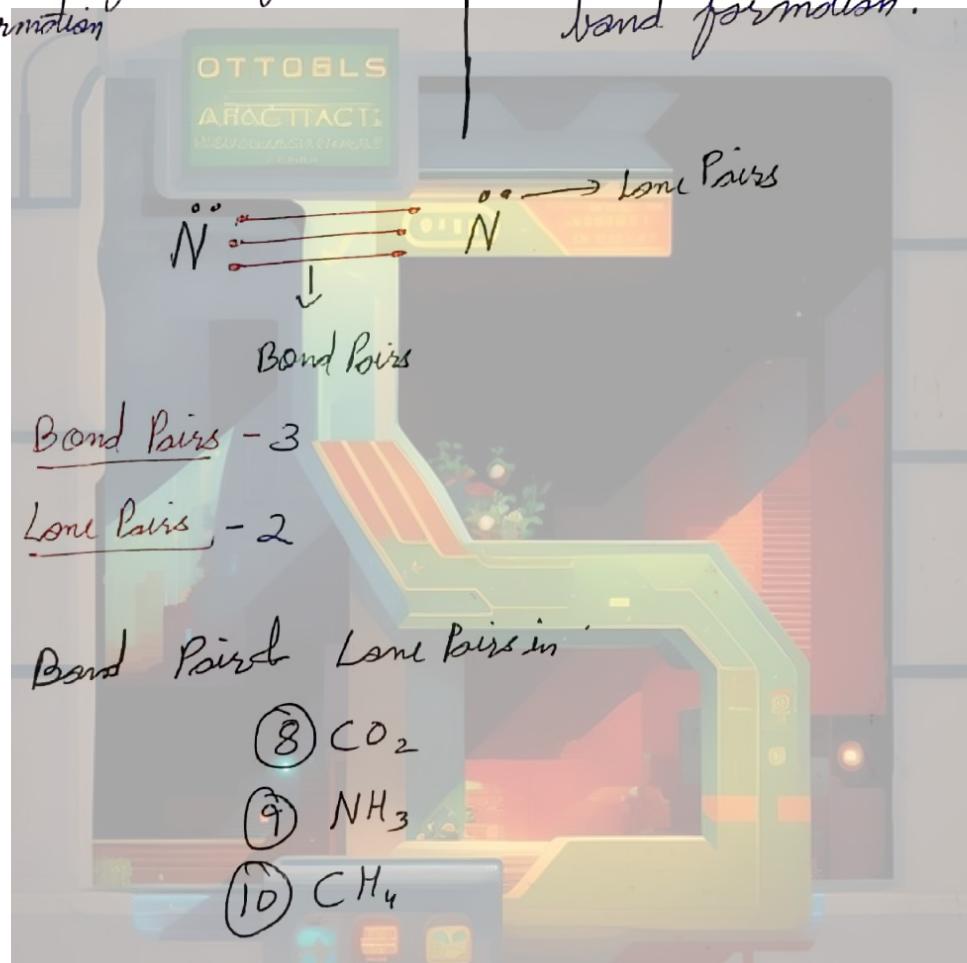
Bond Pair (B.P)

→ Pairs of e^- which are participating during bond formation

Lone Pair (L.P)

→ Pairs of e^- which do not participating during bond formation.

Eg1.



Q Count

- ① Cl_2
- ② H_2O
- ③ CS_2
- ④ SF_2
- ⑤ HCl
- ⑥ NCl_3
- ⑦ C_2F_2

Bond Pairs Lone Pairs in

- ⑧ CO_2
- ⑨ NH_3
- ⑩ CH_4

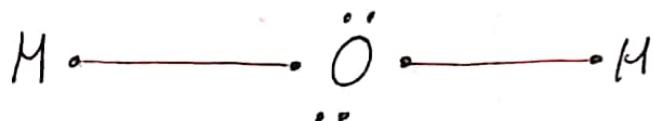
①



Bond Pair $\rightarrow 4$

Lone Pair $\rightarrow 6$

②

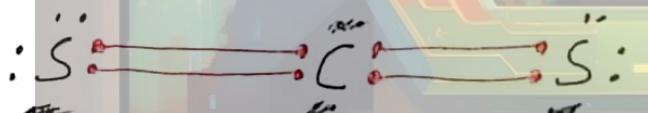


③

BP - 1
LP - 3

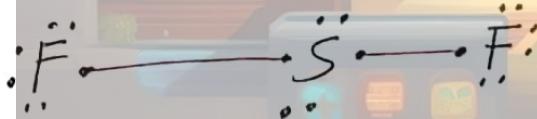


④



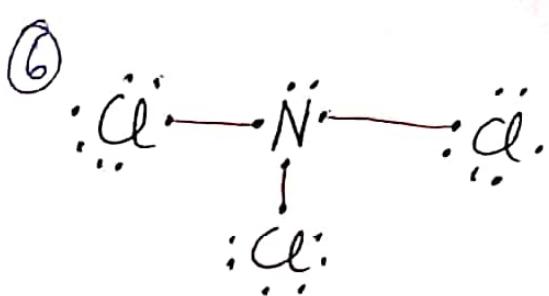
BP - 4
LP - 8

⑤

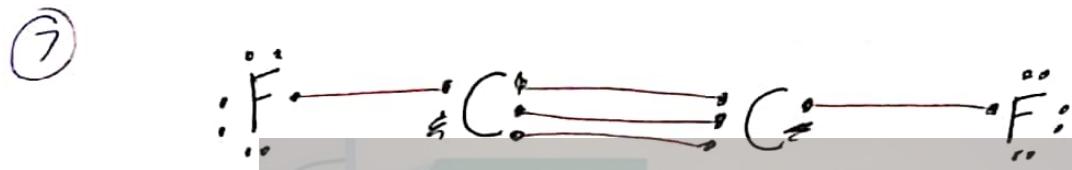


BP - 2
LP - 18

⑥ 180



BP - 3
LP - 10



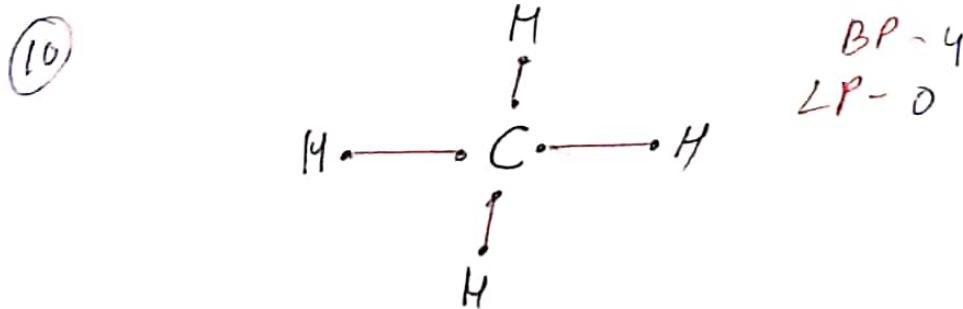
BP - 5
LP - 6



BP - 4
LP - 4



BP - 3
LP - 1



CovaLency - The number of electrons ^{not} participating of an element in p covalent bond formation is known as its covalency.

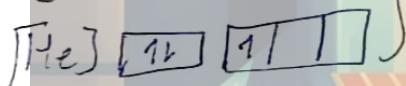
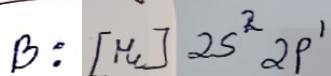
It is defined as the number of covalent bonds formed by an atom in a covalent molecule or compound.

Number of unpaired electrons in outermost shell in ground state or in excited state decides co-valency of an atom.

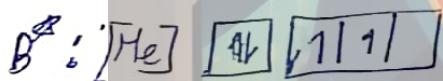
→ In S-block & P-block elements, the number of e⁻ in outermost shell is called valence electron.

→ In D-block elements, outermost 's' & penultimate 'd' electrons are collectively called valence electrons.

Boron Family



$$\text{CovaLency} = 3 \quad (\text{does not exist})$$



$$C.V. = 3$$

$$\text{CovaLency} = 3$$

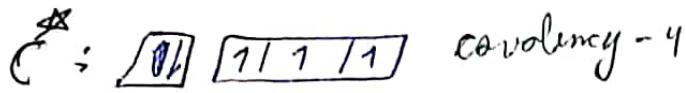
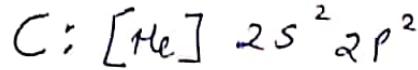
→ The excitation of 1s is not possible since shell changes require very high amount of energy.

→ Boron, Aluminium, Indium, Gallium forms bond only in excited state and their co-valency is always 3.

e.g. ① BCl_3 , BF_3 , $AlCl_3$ (covalent, not ionic bond)

$$\Delta EN < 2.1$$

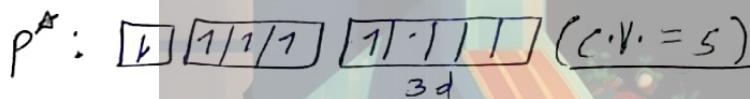
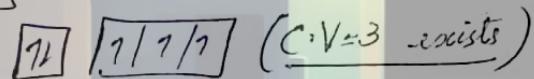
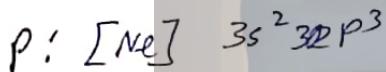
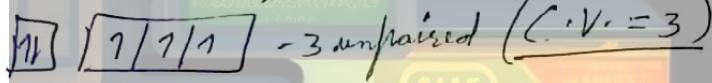
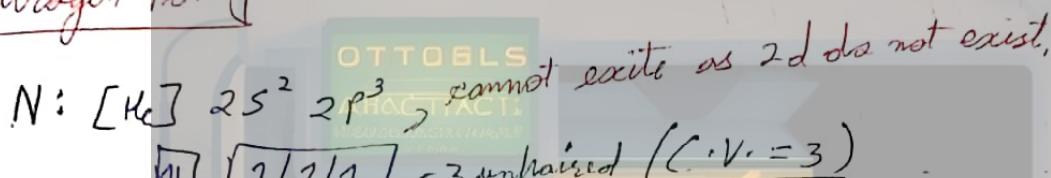
Carbon family - (tetravalent)



Eg ①. CH_4, CO_2, CCl_4

$\rightarrow C, Si, Ge, Sn, Pb$ prefers to form 4 bonds ($C.V. = 4$)

Nitrogen Family



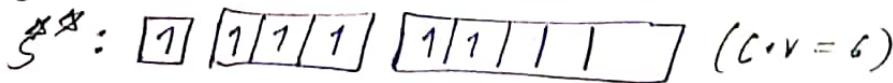
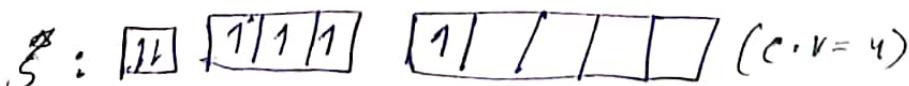
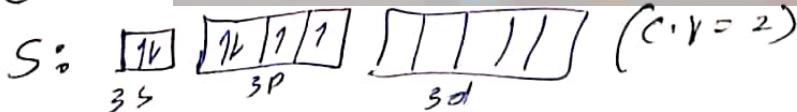
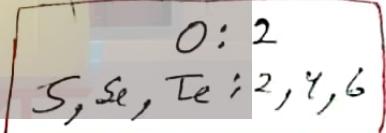
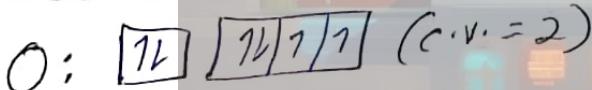
\rightarrow Nitrogen forms strong bonds

\rightarrow Nitrogen has covalency 3.

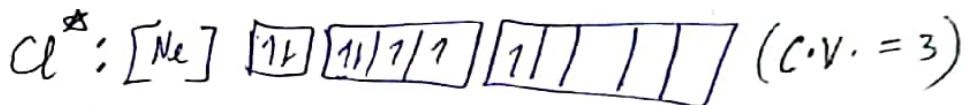
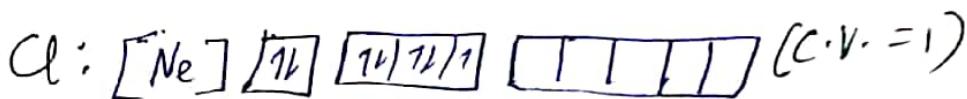
$\rightarrow P, As, Sb$ has covalency both 3 & 5.

Eg ① PCl_3, PCl_5

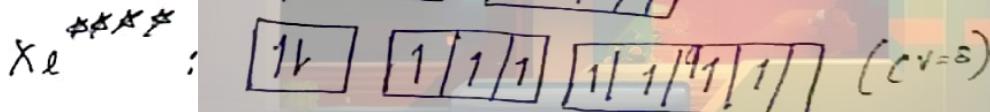
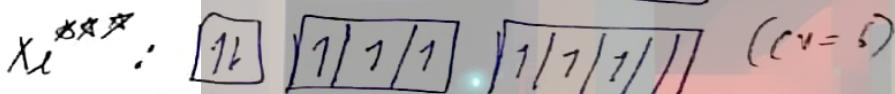
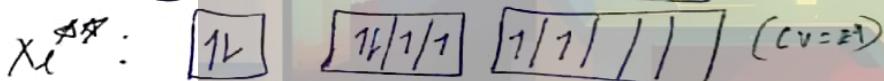
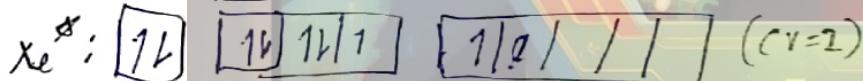
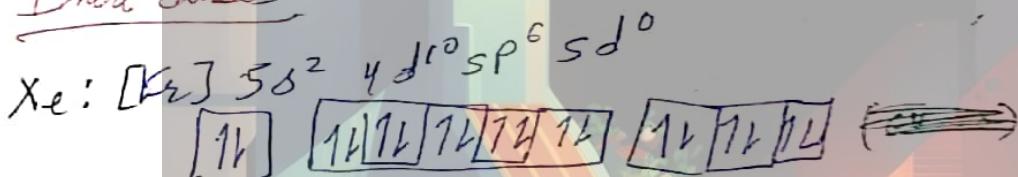
Oxygen family



Molegan Family

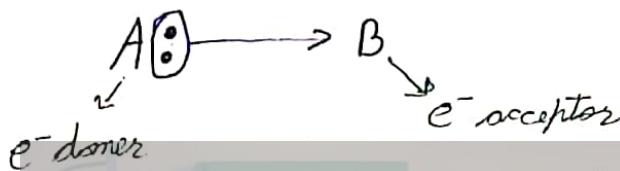


Inert Gases



3. α -coordinate Bond (Donor Bond)

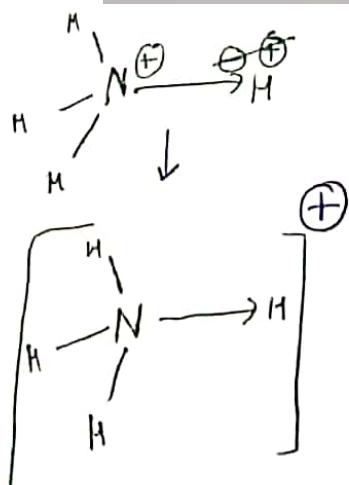
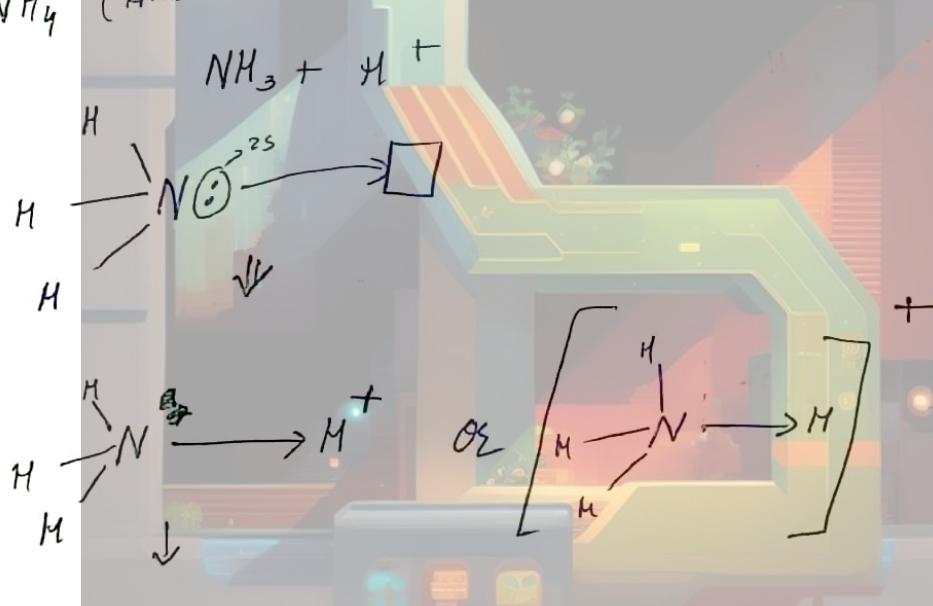
- It is formed by unequal sharing of electrons.
- It means out of two atoms between which a bond is formed, only one atom provides both the electrons that are shared between them.
- It is a type of covalent bond.



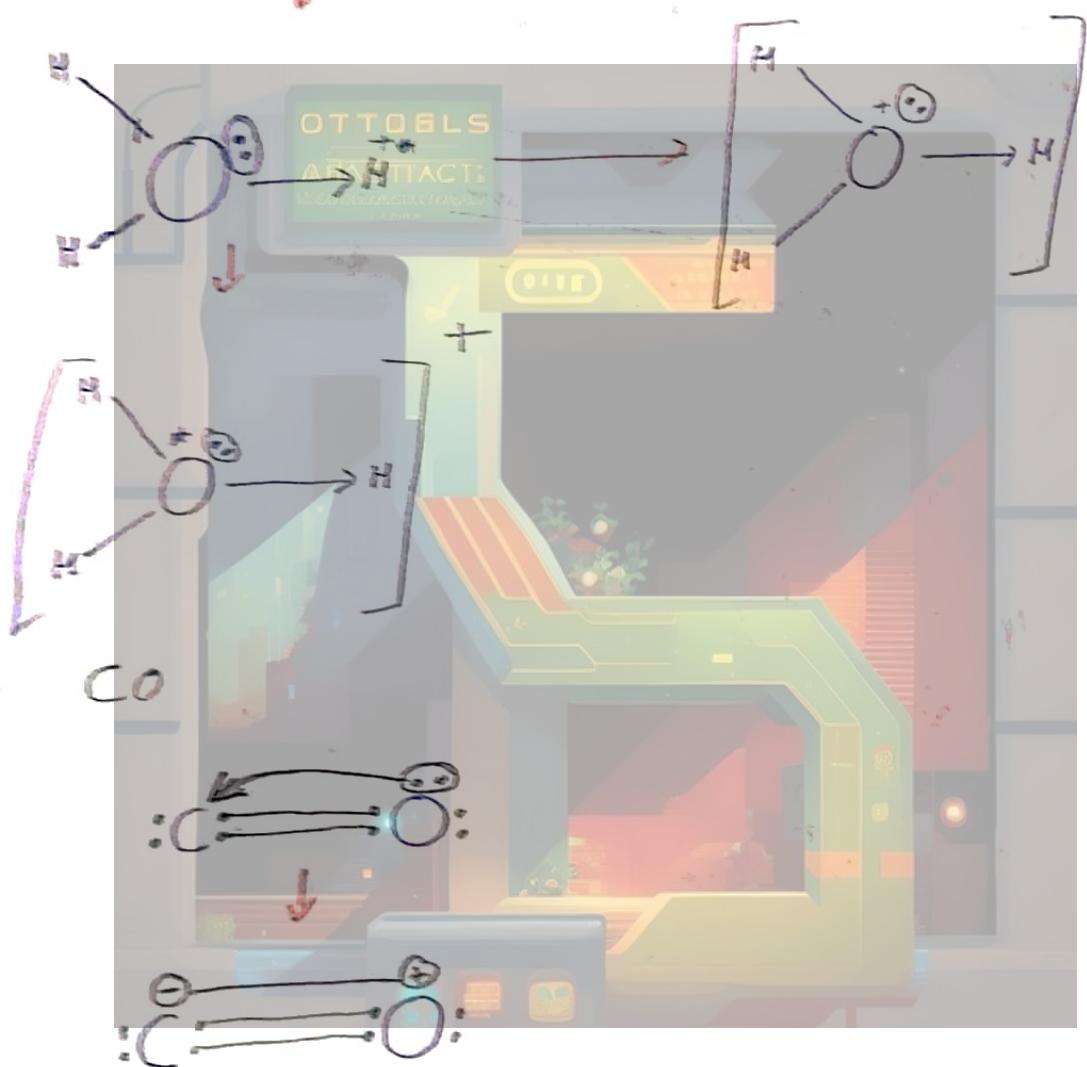
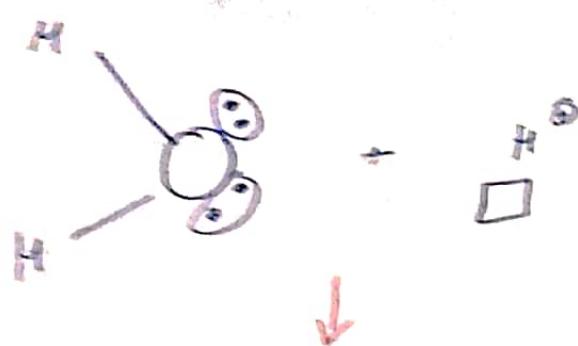
Condition to form α -coordinate bond :-

- ① One atom have at least 1 lone pair.
- ② One atom have atleast one vacant 2s orbital.

Eg NH_4^+ (Ammonium ion)



50 (H_3O^+)



Lewis Dot Structure (L-D-S) :-

- According to Lewis Theory, every atom tries to obtain inert gas configuration for achieving stability.
- He proposed a theory that a chemical bond will be formed by an atom so that it has an octet configuration (except hydrogen - Duplet)

Eg ① HF



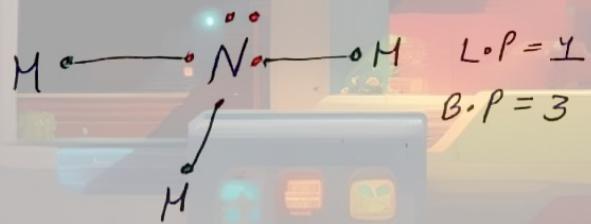
Rules to draw LDS

- ① Close center Central atom -
 → Hydrogen & Fluorine can never be central atom
 → Usually central atom is least electron negative atom and also lesser in number

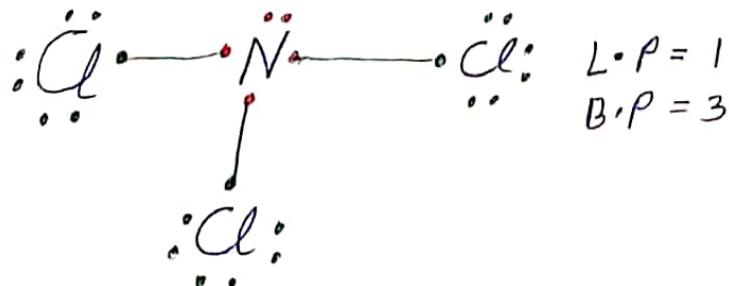
① CO_2



② NH_3



③ NCl_3



④ HCN (Cyanide)

⑤ PH₃

⑥ HNC (Iso-Cyanide)

⑦ H₂S

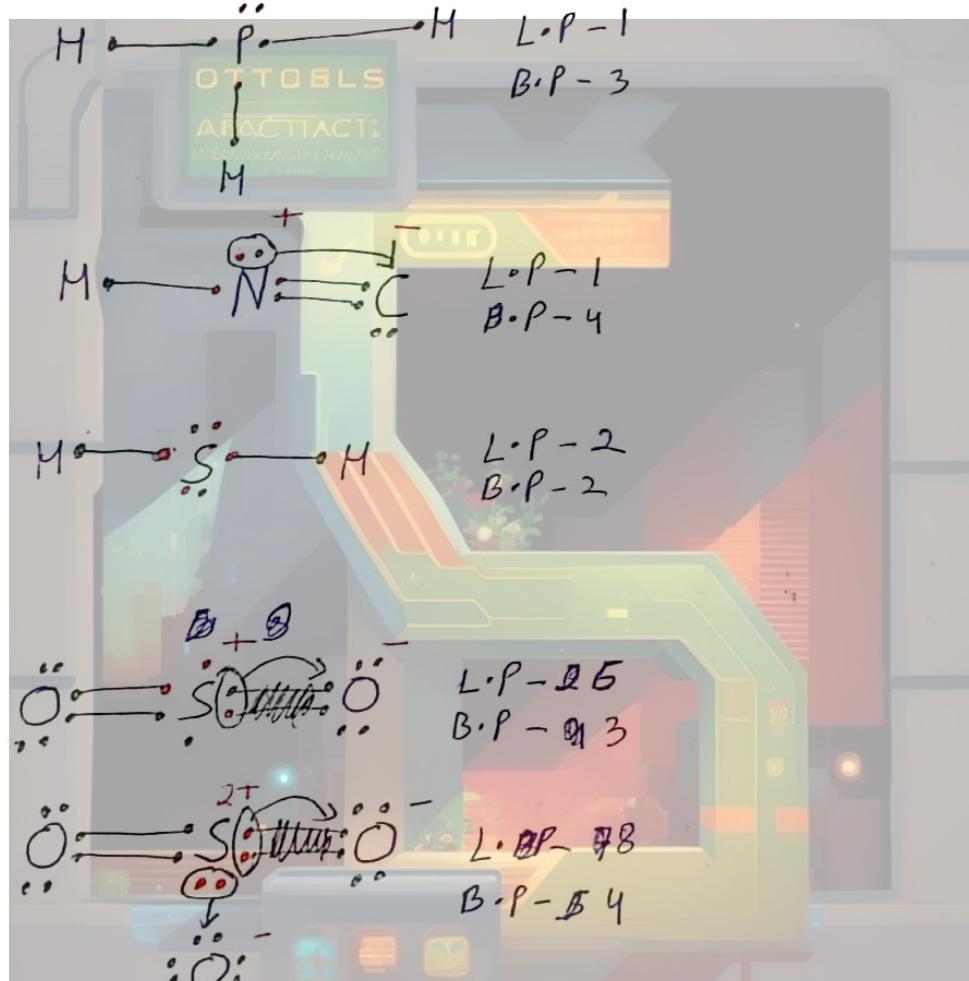
⑧ SO₂

⑨ SO₃

⑩



⑤



O:

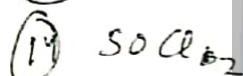
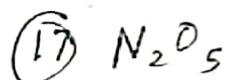
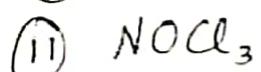
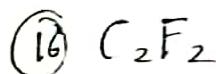
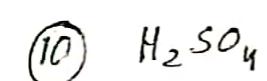
11	11	11	1	1
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O:

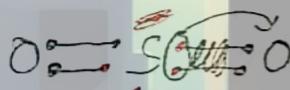
11	11	11	11
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 → To create vacant orbital for co-ordinate bond.

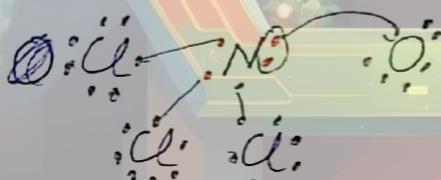
Note - ① Oxygen atom in ground state cannot form coordinate bond because it has no vacant orbital.
 ② However if 1 e⁻ of 2p subshell gets paired in another orbital of 2p subshell then a vacant orbital is created.



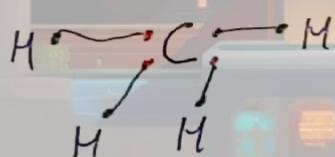
⑩



⑪



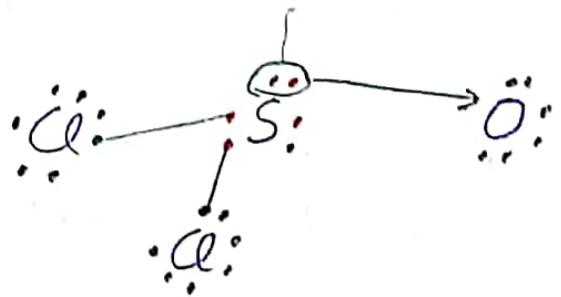
⑫



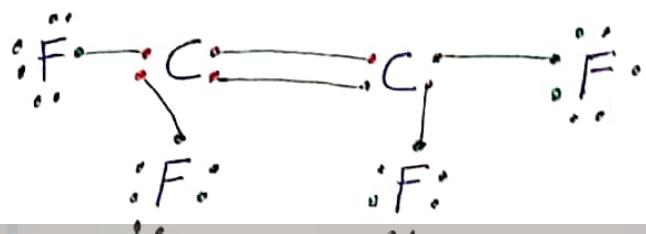
⑬



(14)



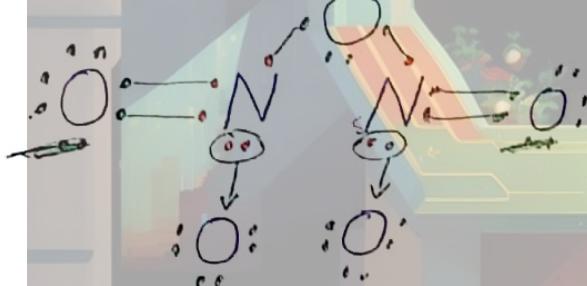
(15)



(16)



(17)

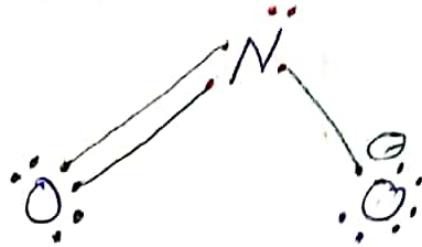


LDS of charged species

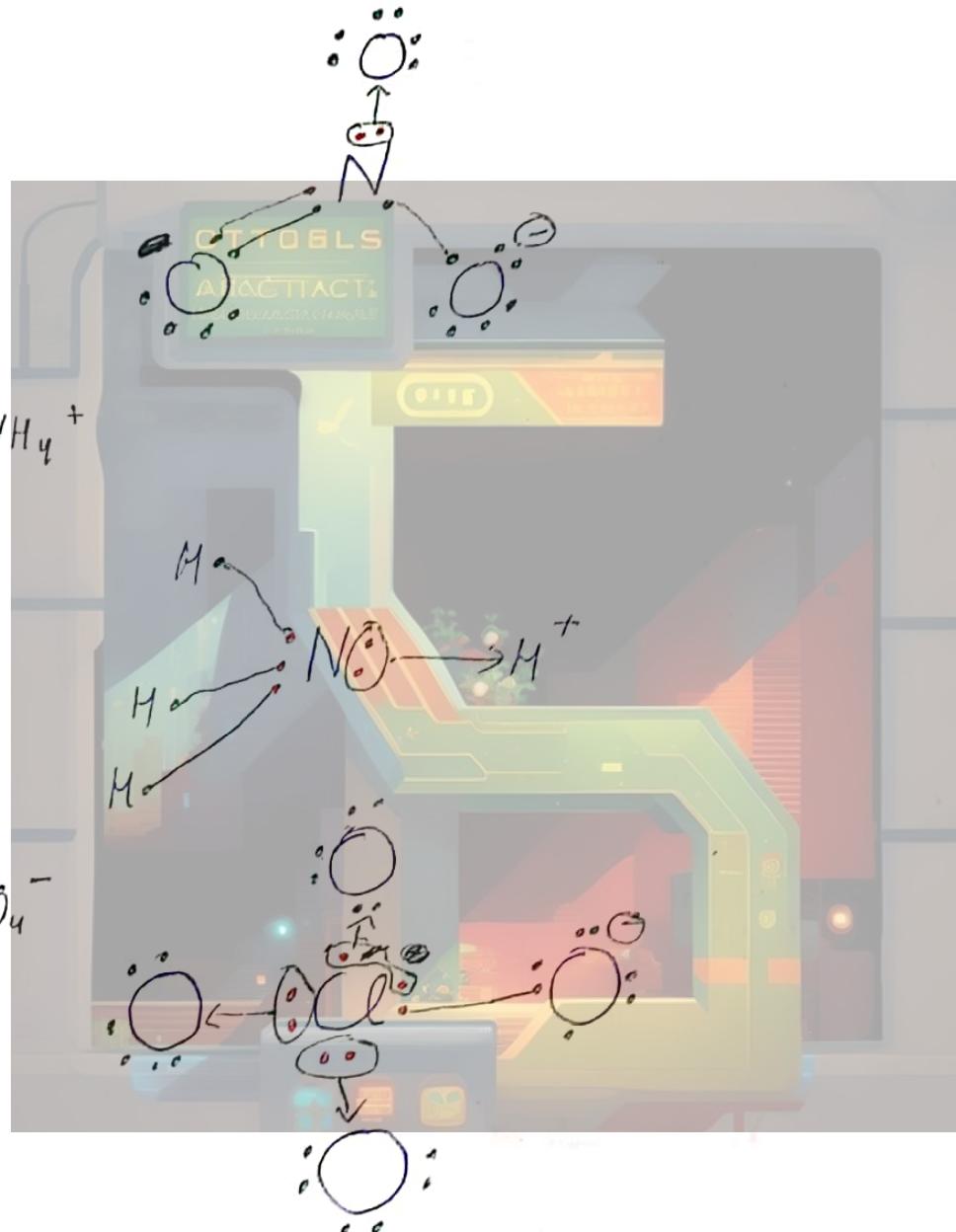
- \ominus ve charge carried by more EN atom
- \oplus ve charge carried by less EN atom.

(190)

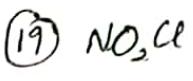
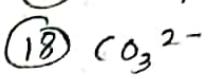
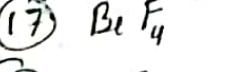
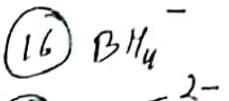
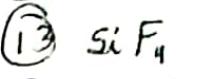
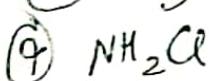
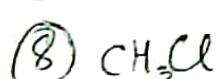
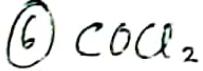
Eg ①. NO_2^-



② NO_3^-

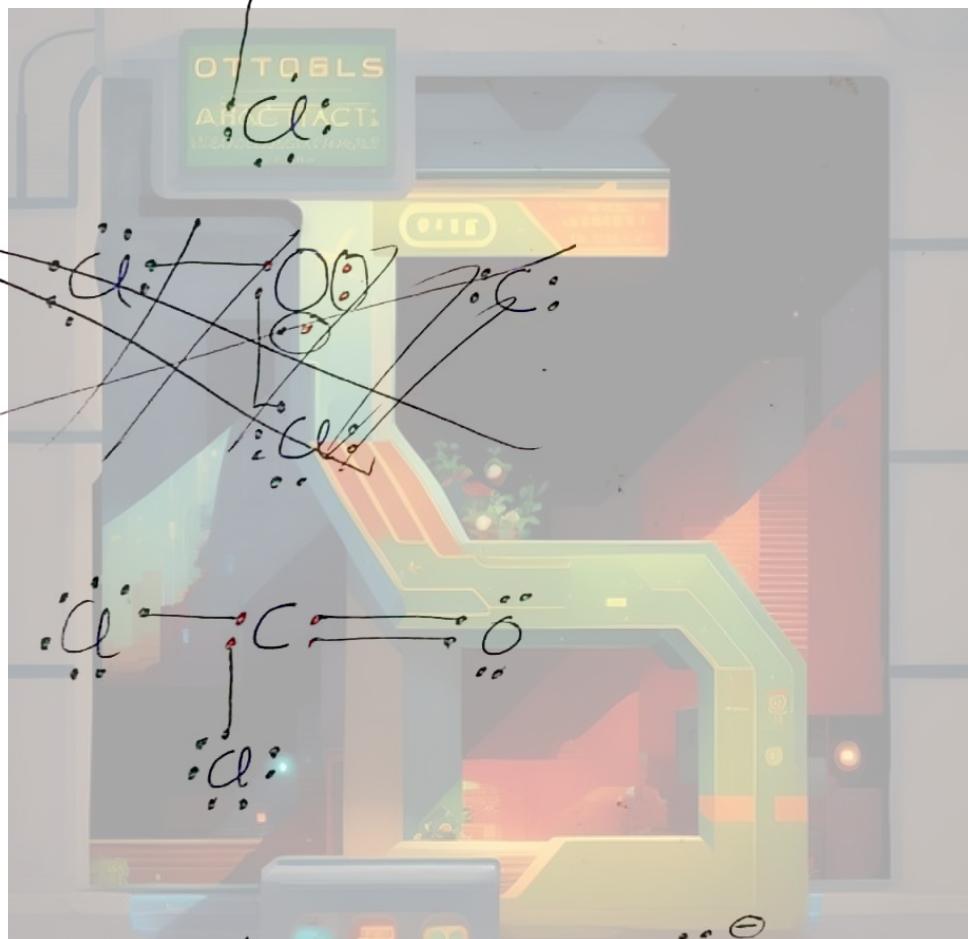


③. NH_4^+



(Tildenireagenz)

(5)



(6)

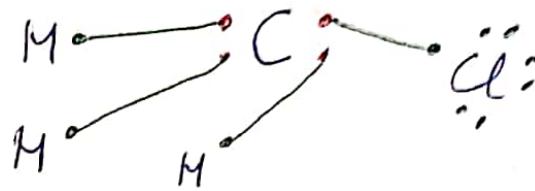


or



(192)

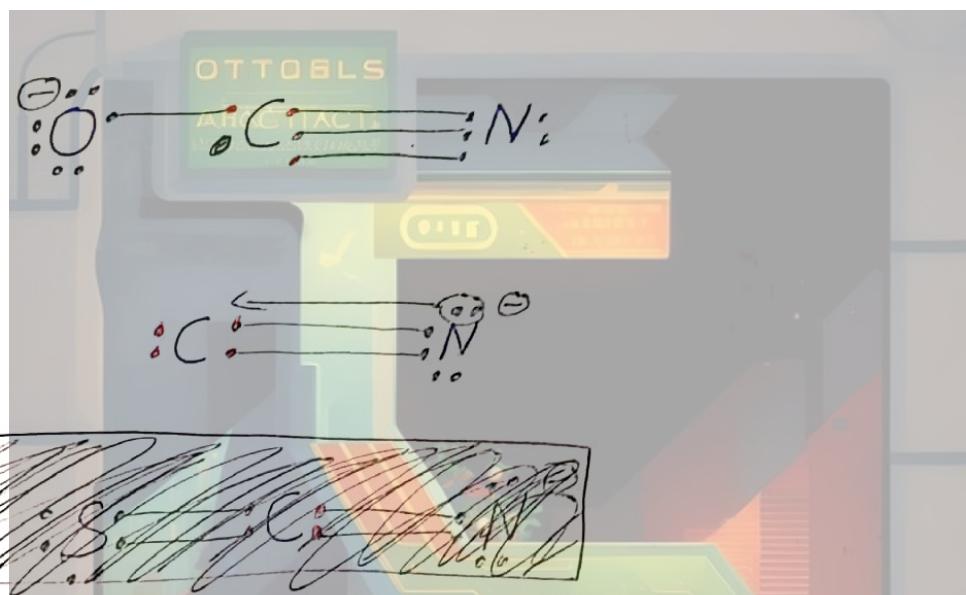
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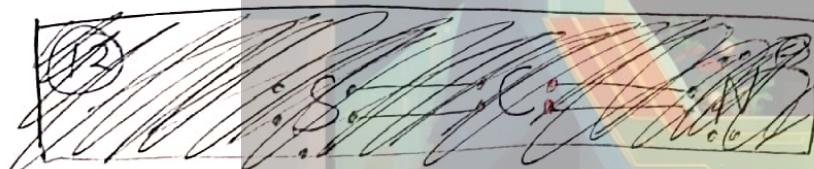
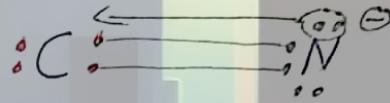
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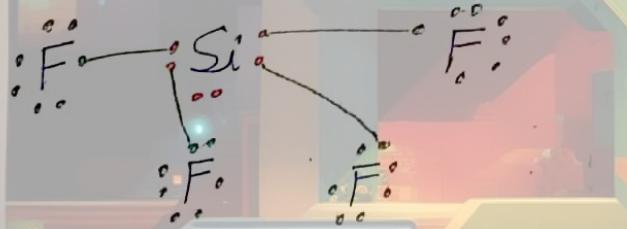
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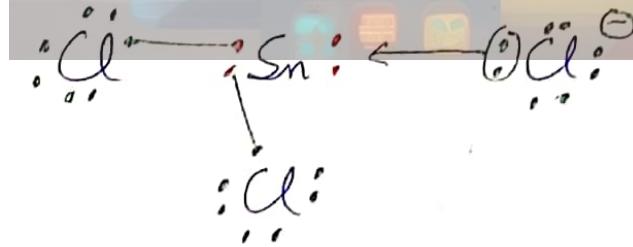
(11)



(12)

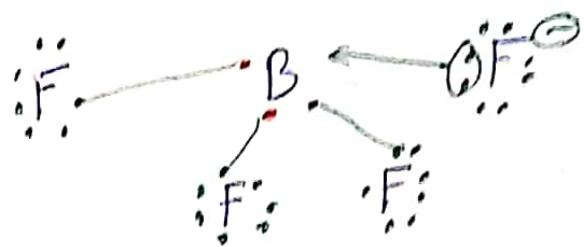


(13)

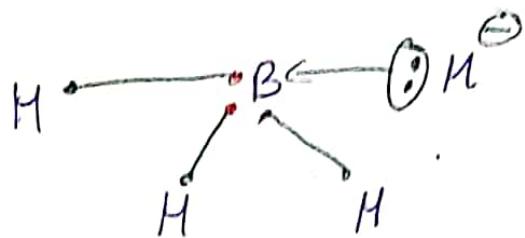


(14)

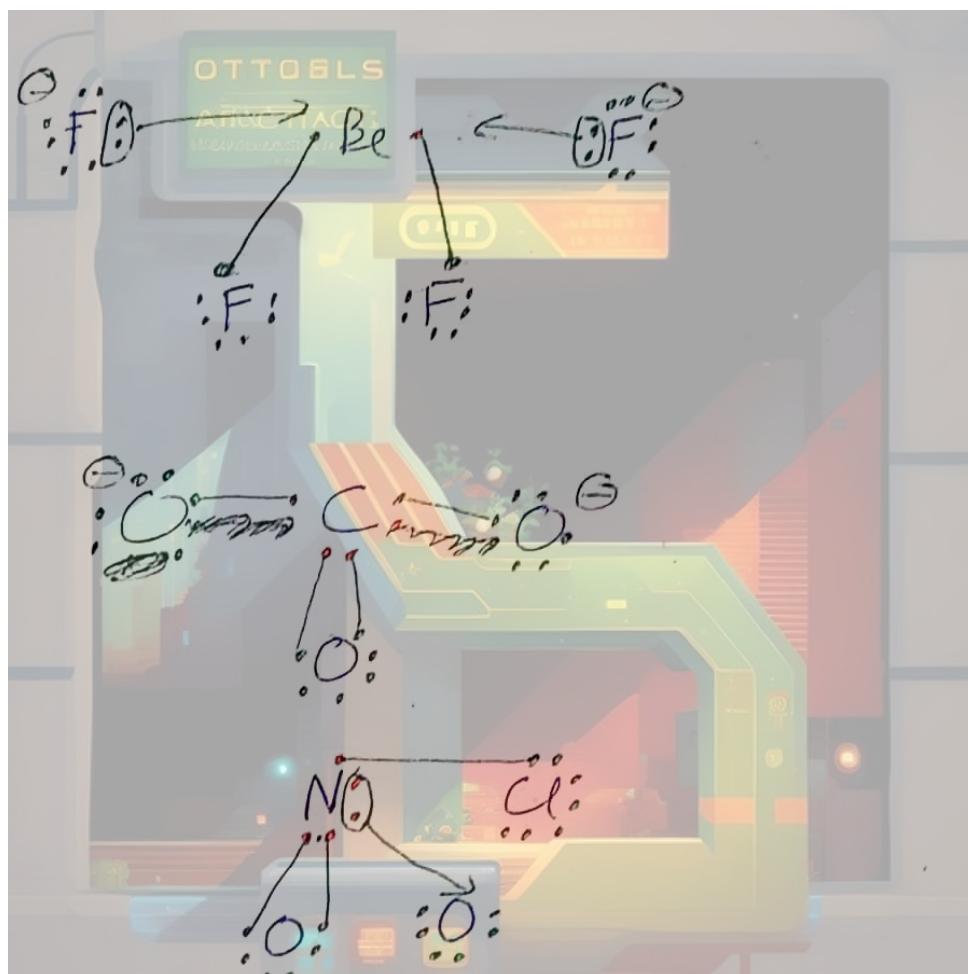
(15)



(16)



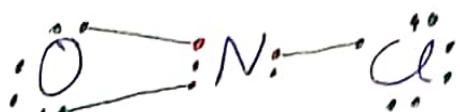
(17)



(18)

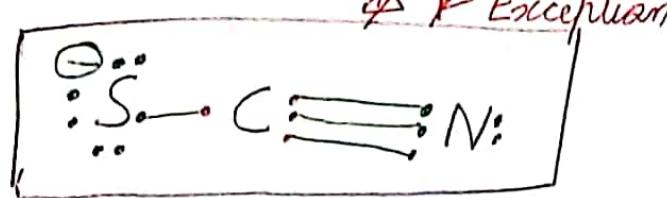


(19)

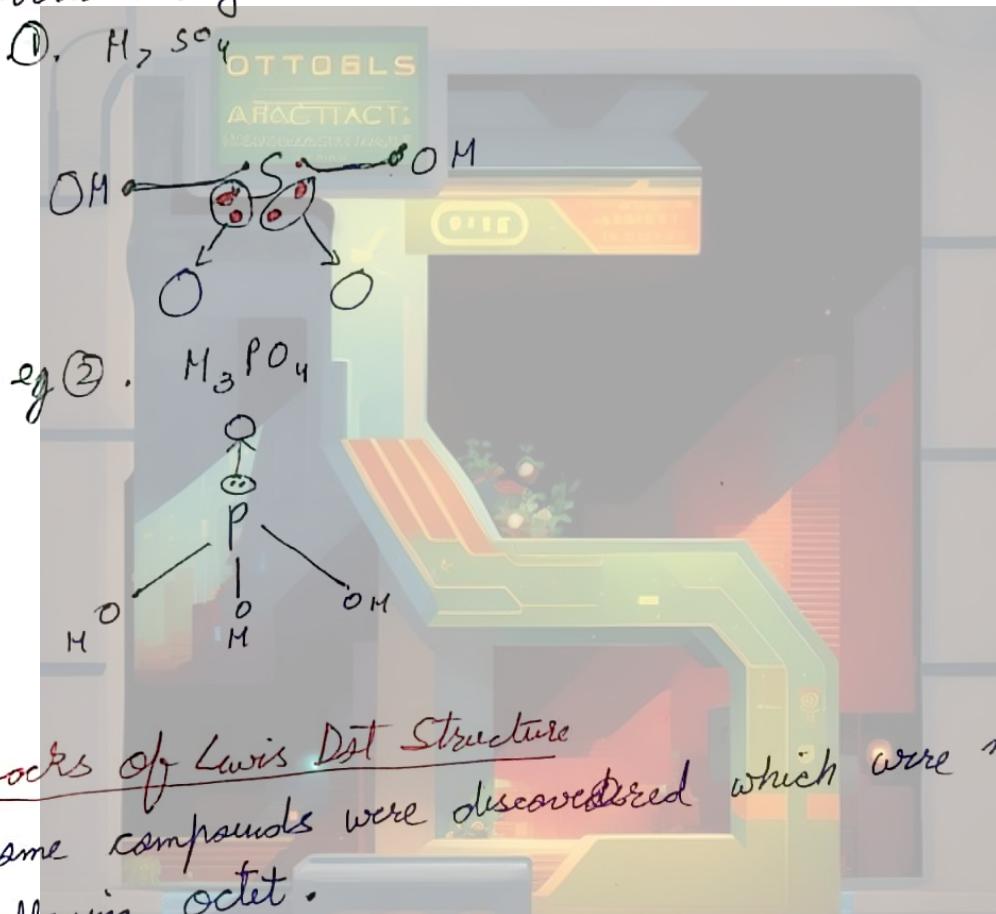


(194)

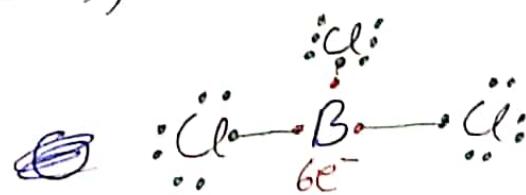
(12)

Note :-

- ① If oxygen and hydrogen both are present together in a molecule. Then connect them with central atom in the form of OH bond. (few exceptions are there will be covered in oxy acids).

eg ①. H_2SO_4 Drawbacks of Lewis Dot Structure

- ① Some compounds were discovered which were not following octet.
- There are 3 categories
- ① hypervalent - Species in which central atom occupy less than 8 valent electrons after bonding.

eg. BF_3 , BCl_3 , BBr_3 , BT_3 , AlCl_3 

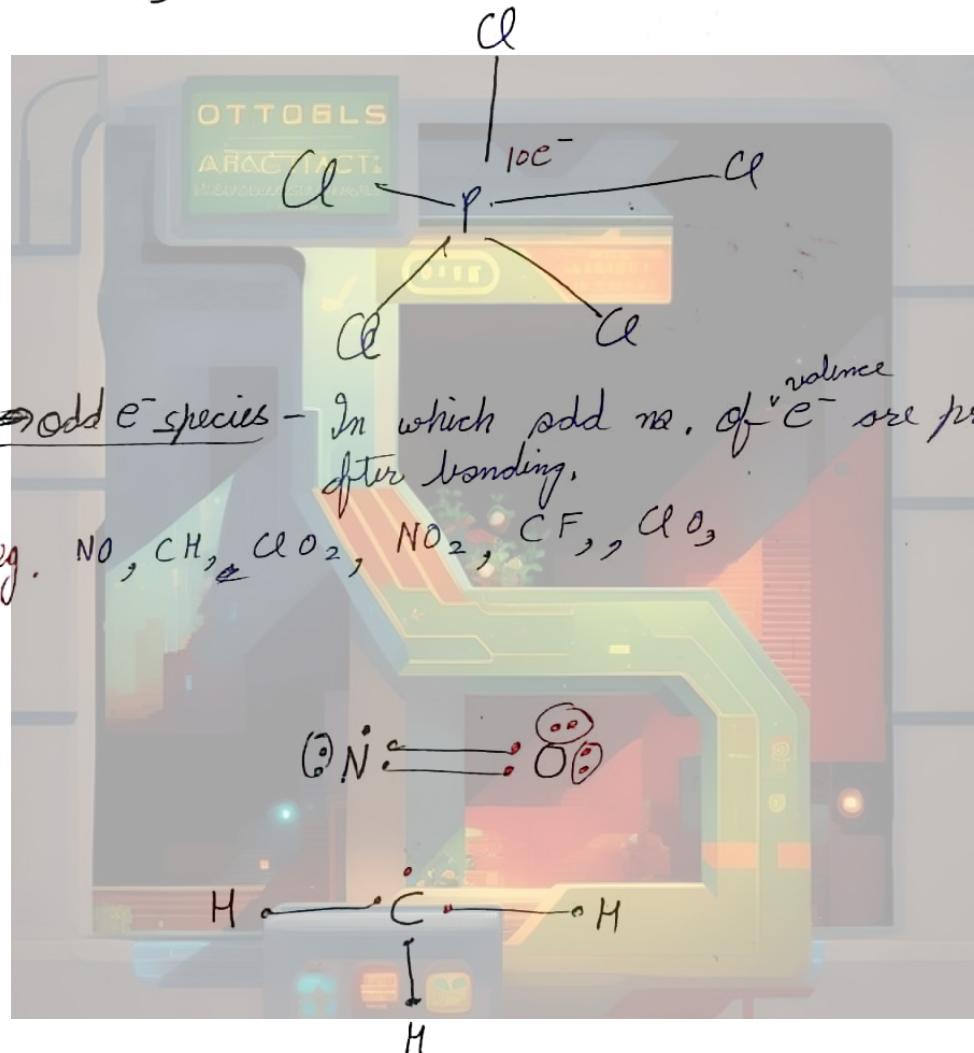
(195)

→ AlF_3 is not hypervalent because it has complete octet due to ionic nature.

→ AlF_3 is ionic but AlCl_3 , AlBr_3 & AlI_3 are covalent compound. (SEN)

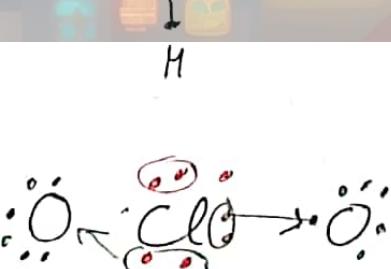
② Higher Valent - Central atom occupies more than 8 valence e⁻ after bonding.

e.g. PCl_5 , XeE_6 , XeF_4 , XeF_6 , IF_7 , IF_5



③ Odd e⁻ species - In which odd no. of "e⁻" are present after bonding.

e.g. NO , CH_3 , ClO_2 , NO_2 , CF_3 , ClO_3



② Lewis theory cannot explain shape of or geometry of molecule.