Chemistry - 1

Ch1 Quantum Numbers & Electronic Configration -

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Chemistry

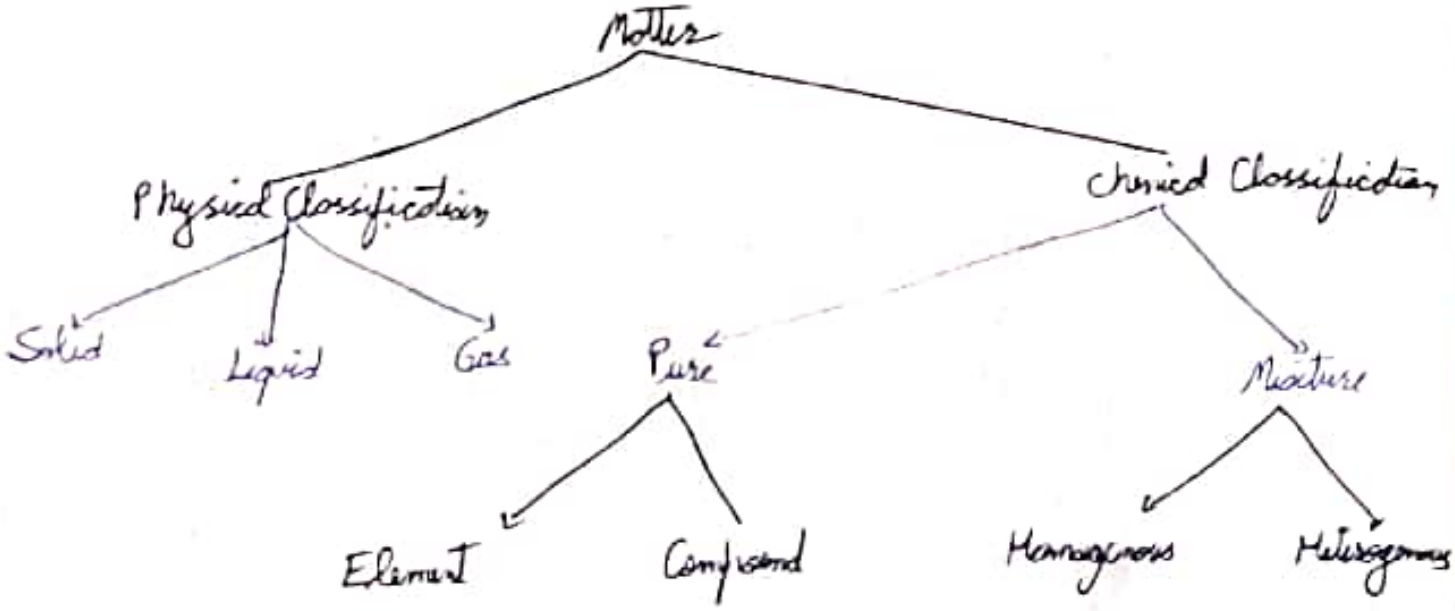
Inorganic | Organic | Physical

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

Inorganic

Quantum Numbers & Electronic configuration

Matter → Anything which has mass & occupies some space



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

Compound → Made up of two or more than two types of atoms
eg 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

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

Heterogeneous Mixture → It is a mixture of two or more types of substances having non-uniform composition.
eg Oil + water
Sand + water
Smoke, 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
		Bronze	

Atom → ① Smallest unit of matter
② It may or not exist in free state in nature
eg N_2, Cl_2
③ It is electrically neutral
④ Made up of 3 fundamental particles - electrons
- protons
- neutrons

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Particle	Charge	Mass	Nature of charge	Position
1. Electron $\boxed{{}^{0}_{-1}e}$	-1.602×10^{-19} C or -4.8×10^{-10} e.s.u.	9.11×10^{-28} g	Negative	Outside the Nucleus
2. Proton $\boxed{{}^{1}_{1}p}$	$+1.602 \times 10^{-19}$ C or $+4.8 \times 10^{-10}$ e.s.u.	1.6725×10^{-24} g or $\frac{1}{1837}$ mass of e^-	Positive	Inside the Nucleus
3. Neutron $\boxed{{}^{1}_{0}n}$	0	1.675×10^{-24} g	Neutral	Inside the nucleus

e.s.u - electrostatic unit of charge

$\boxed{1.67 \times 10^{-24} \text{ gram} = 1 \text{ amu}}$

Representation of a atom

$${}^{A}_{Z}X \quad \text{or} \quad {}_{Z}X^{A}$$

A - Atomic mass number

Z - Atomic number

no. of p = Z

no. of e = p = Z (In a neutral absence of charge)

no. of e = Z - charge (for cation)

Z + charge (for anion)

atomic mass (A) = n + p

eg. ${}^{16}_{8}O$

O - Symbol of oxygen atom

16 - A - mass

8 - Z - atomic no

p = Z = 8

e = p = 8

n = A - Z = 16 - 8

= 8

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Ex - 1

CH_4 $p = 6 + 4 = 10$
$e = 6 + 4 = 10$
$n = (12 - 6) + 4(1 - 1) = 6 + 0 = 6$

H_2O $p = 2 + 8 = 10$
$e = 2 + 8 = 10$
$n = 0 + 8 = 8$

H_2O_2 $p = 2 + 16 = 18$
$e = 2 + 16 = 18$
$n = 0 + 16 = 16$

N_2O $p = 2 + 8 = 10$
$e = 2 + 8 = 10$
$n = 2 + 8 = 10$

NH_3 $p = 7 + 3 = 10$
$e = 7 + 3 = 10$
$n = 7 + 0 = 7$

CO_3^{2-} $p = 6 + 24 = 30$
$e = 6 + 24 + 2 = 32$
$n = 6 + 24 = 30$

SO_4^{2-} $p = 16 + 32 = 48$
$e = 16 + 32 + 2 = 50$
$n = 16 + 32 = 48$

O_2 $p = 16$
$e = 16$
$n = 16$

$+1$ or 1^+	→ monovalent cation
$+2$ or 2^+	→ Divalent cation
$+3$ or 3^+	→ Trivalent cation
-3 or 3^-	→ Trivalent anion
-2 or 2^-	→ Divalent anion
-1 or 1^-	→ monovalent anion

1 - mono	6 - Hexa
2 - Di	7 - Hepta
3 - Tri	8 - Octa
4 - tetra	9 - nona
5 - penta	10 - deca

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

Atomicity → Total number of atoms present in a molecule is called its atomicity.

eg. Ne → Monoatomic (atomicity -1)
O_2 → Diatomic (atomicity - 2)
O_3 → Triatomic (atomicity - 3)
P_4 → Tetratomic (atomicity -4)
S_8 → Octatomic (atomicity - 8)

Isotopes

1. **Isotopes** - Atoms having same atomic number (Z) of same element but different mass number (A).

eg. 1 $^{1}_{1}H$ → Protium (H), $^{2}_{1}H$ → Deutrium (D), $^{3}_{1}H$ → Tritium (T)

eg 2 $^{35}_{17}Cl$ → 75%, $^{37}_{17}Cl$

Average A = 35.5

eg. 3 $^{12}_{6}C$ → 98.9, $^{13}_{6}C$, $^{14}_{6}C$

☆ Different number of neutrons.

2. **Isobars** - Atoms having same mass number but different atomic number. Atoms of different elements.

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eg 1. $^{14}_{6}C$ $^{14}_{7}N$

eg 2. $^{40}_{18}Ar$ $^{40}_{20}Ca$

3. Isotones: Species having same number of neutrons.

eg 1.

$^{3}_{1}H$	$^{4}_{2}He$
$n = 3-1$	$n = 4-2$
$= 2$	$n = 2$

eg 2.

$^{23}_{11}Na$	$^{24}_{12}Mg$
$n = 23-11$	$n = 24-12$
$= 12$	$= 12$

4. Isodiaphers - Atoms with same value of $(n-p)$ or $(A-Z)$

eg. 1

$^{19}_{9}F$	$^{23}_{11}Na$
$A-Z$	$A-Z$
$19-18$	$23-22$
1	1

$^{238}_{92}U$	$^{234}_{90}Th$
$A-Z = 54$	$A-Z = 54$

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5. <u>Iso-Electronic</u> – Species (atom/ion/molecule) having same number of total electrons.

eg 1. $^{20}_{10}Ne$ | $^{23}_{11}Na^{+}$

$e = 10$ | $e = 11 - 1 = 10$

eg 2. $_{6}C^{4-}$ $_{7}N^{3-}$ $_{8}O^{2-}$ $_{9}F^{1-}$ $_{10}Ne$ $_{11}Na^{+}$

$_{12}Mg^{2+}$ $_{13}Al^{3+}$ [All have $10e^-$]

eg 3. BO_3^{3-} CO_3^{2-} NO_3^- [All have $32e^-$]

eg 4. SO_4^{2-} PO_4^{3-} ClO_4^- (All have $50e^-$)

eg 5. H_2S HCl Ar (All have $18e^-$)

★ All iso-topes are iso-electronic.

6. <u>Isosters</u> – Species having same number of atoms as well as same no. of electrons.

eg. N_2O | CO_2

atoms = 3 | atoms 3

$e^- = 22$ | $e^- = 22$

★ All Isosters are Iso electronic but not all Iso-electronics are Isosters.

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H.W.

Ex. O-1 (Q1-Q4)

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. a) b) d)

S-2

Q6-C)

Q7-b)

JEE-Mains

Q1. 1)

Q6 1)

Q7 4)

Q9 4)

Q11. 2)

Quantum Numbers

atom — ~~Principal Quantum Number (n)~~

↓

Principal Quantum Number (n) — Shell/orbit — ~~Azimuthal Quantum Number (l)~~

↓

Azimuthal Quantum Number (l) — Subshell — ~~Magnetic Quantum Number (m or m_l)~~

↓

Magnetic Quantum Number (m or m_l) — Orbital — ~~Spin Quantum Number (s, m_s)~~

↓

Spin Quantum Number (s or m_s) — Present/Absent

☆ First 3 Quantum Numbers gives information about the orbitals

☆ First 4 Quantum Numbers gives information about the electrons

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

1. **Principle Quantum Number (n)** - Principal Quantum Number 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 to ∞

$$\begin{array}{ccccccc} 1 & 2 & 3 & 4 & 5 & 6 & \ldots \infty \\ \downarrow & \downarrow & \downarrow & \downarrow & \downarrow & \downarrow & \\ K & L & M & N & O & P & \end{array}$$

☆ Maximum number of possible shells in an atom is infinite

n ↓ Attraction by Nucleus ↑ Stability ↑ Energy ↓

☆ n decides the energy of a 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 $2n^2$

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Shell no.	Max no of orbitals	Max no. of electrons
$n = 1$ (K)	$= (1)^2 = 1$	$= 2(1)^2 = 2$
$n = 2$ (L)	$= (2)^2 = 4$	$= 2(2)^2 = 8$
$n = 3$ (M)	$= (3)^2 = 9$	$= 2(3)^2 = 18$
$n = 4$ (N)	$= (4)^2 = 16$	$= 2(4)^2 = 32$
$n = 5$ (O)	$= (5)^2 = 25$	$= 2(5)^2 = 50$

2. Azimuthal Quantum Number (l) / secondary Q. No. / subsidary 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) ~~and~~ name of subshell
   b) energy of subshell (lT, $\propto$ r)
   c) shape of subshell

Value of l	Name of Subshell	Shape of subshell
$l = 0$ (s)	Sharp	Spherical
$l = 1$ (p)	Principal	Dumbell
$l = 2$ (d)	Diffused	Double Dumbell
$l = 3$ (f)	fundamental	complex (not in JEE)
$l = 4$ (g)	—	—

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☆ 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 -

$$[ns < np < nd < nf]$$

eg. $4s < 4p < 4d < 4f$

3. <u>Magnetic Quantum Number</u> (m or m_l)

1. It gives the information about possible orientation (arrangement) of an ~~orb~~ orbital in space.
2. Values of m depends upon the value of l and ranges from $-l$ to $+l$ (including 0) in integral steps.

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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)
J-M
Q5. 1)
Q8. 2)
Q10. 2)
Q13 A)
Q14 3)
Q15 C)
J-A
Q1. 9
Q2. 6

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n	l	m	no. of orbitals	no. of e^-
$n=1$	$l=0(s)$	$m=0$	$1=(1)^2$	$1\times2=2$
$n=2$	$l=1(p)$	$m=-1,0,+1$	$(2)^2=4$	$4\times2=8$
$n=3$	$l=2(d)$	$m=-2,-1,0,1,2$	$(3)^2=9$	$9\times2=18$
$n=4$	$l=3(f)$	$m=-3,-2,-1,0,1,2,3$	$(4)^2=16$	$16\times2=32$

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

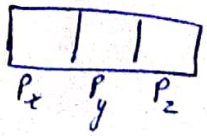
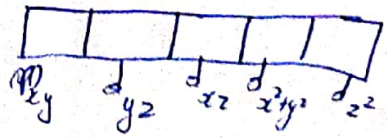
eg. $l=0(s)$

~~$n=0$~~ $2\times0+1=1$ orbital

~~1 orb~~

$l=1(p)$

$2\times1+1=3$ orbitals

l	m	representation of orbital
$l=0(s)$	$m=0$	1
$l=1(p)$	$m=-1,0,1$	3 
$l=2(d)$	$m=-2,-1,0,1,2$	5 

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

✓a) s b) p c) d d) f

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4. ~~Magnetic~~ Spin Quantum Number. (s, m_s)

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

	Clockwise	Anticlockwise
spin up	↑ or ↑	↓ or ↓
spin down	↓ or ↓	↑ or ↑
	$+\frac{1}{2}$	$-\frac{1}{2}$
	$-\frac{1}{2}$	$+\frac{1}{2}$

↑↑ ⇒ Parallel Spin

↑↓ or ↑↓ ⇒ Opposite Spin

Summary of Quantum Numbers

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

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

$n = 3$
$l = 0$ to $(n-1)$
$l = 0, 1, 2$

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

$l = 1$
$m = -1, 0, +1$

Q3. What is the minimum value of m for d-subshell? $-l$ 2

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

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

$l = 0$
$e^- = 6(l+1)^2$
$= 6(1)^2$
$\boxed{= 2}$

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

$l = 1$
$e^- = (2l+1)^2$
$\boxed{= 6}$

Q6. Select correct set of Quantum Nos.

	n	l	m	s
(A)	3	3	-2	$-1/2$
(B)	2	1	-2	$+1/2$
✓(C)	3	2	1	$+1/2$
(D)	4	2	-2	+1

Q7. How many max e^- can be described by Q. Nos. $n = 5$, $l = 2$ in a atom.

$l = 2$
$m = -2, -1, 0, 1, 2$
no. of $e^- = 2(2 \times 2 + 1)$
$\boxed{= 10}$

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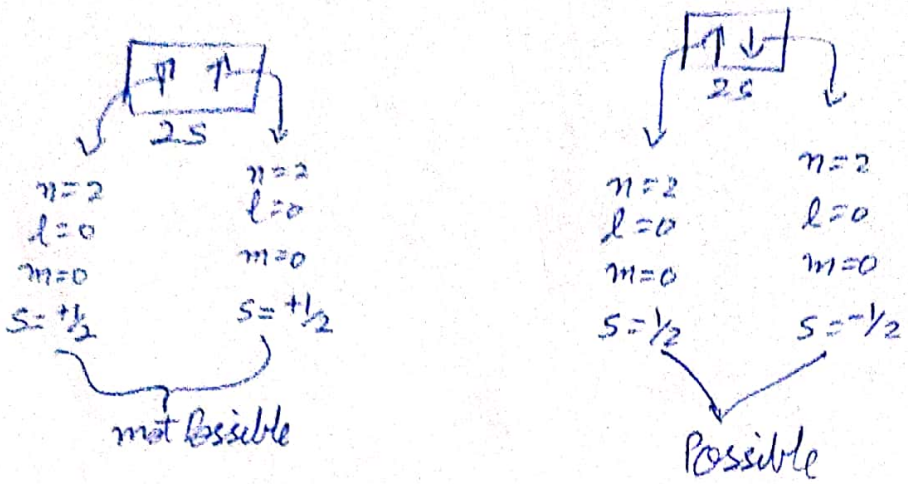
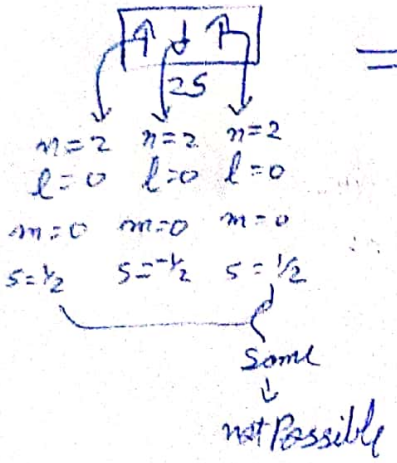
Rule for Filling electrons

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.





⟹ This is the reason, one orbital has only 2 e^-

2. **Afbau'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 ~~level~~ subshell first and once they are filled, then filling of next higher energy subshell occurs.

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

Ex-1

4s	3d
n=4	n=3
l=0	l=2
n+l=4	n+l=5

3d > 4s (energy)

Ex-2

5p	4f
n=5	n=4
l=1	l=3
n+l=6	n+l=7

4f > 5p (energy)

Ex-3

4p	5s
n=4	n=5
l=1	l=0
l+n=5	l+n=5

5s > 4p (energy)

* If two subshells having <u>same value of (n+l)</u>, then subshell with <u>higher value of n</u> has higher energy.

Q energy order?

1. 3s	2p	3d	4f
n=3	n=2	n=3	n=4
l=0	l=1	l=2	l=4
n+l=3	n+l=3	n+l=5	n+l=7

$\boxed{4f > 3d > 3s > 2p}$

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

$ns < (n-3)g < (n-2)f < (n-1)d < (np)$

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ii)

4f	5d	6s	4p
7	7	6	5

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

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

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

$5f > 5d > 4f > 5p > 4d > 5s > 4p > 4s > 3p > 3s > 2p > 2s$

$> 1s$

n/ml	0	1	2	3
1	1s			
2	2s	2p		
3	3s	3p	3d	
4	4s	4p	4d	4f
5	5s	5p	5d	5f
6	6s	6p	6d	
7	7s	7p		

Q

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Q1. $4d_{xy}$ $4f$ $3d_{yz}$ $3d_{z^2}$ $5p_y$

6 7 5 5 ~~4~~ 6

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

<u>degenerate orbitals</u> – All the orbitals of a particular ~~sub~~ subshell have same energy.

→ The orbitals having same value of n & l but different value of m have same energy in the absence of a electric & magnetic field.

→ These orbitals having same energy of a particular subshell are known as ~~dege~~ degenerate orbitals.

<u>Degeneracy</u> :- Total no. of orbitals with same energy.

Q. find degeneracy

i) p subshell - 3

ii) d subshell - 5

iii) f subshell - 7

iv) s subshell - Not-Defined (D = 0)

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

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H. W. 09-4-24

O-1 (Q5, 7, 15)
O-2 (Q2, Q5)
JM (Q5, 8, 10, 13, 14, 15)
JA (Q1, 2)

Answers:-

O-1
Q5 - B)
Q7 C)
Q15. D)

O-2
Q-2 C)
Q5. B, D

JM
Q5 1)
Q8 2)
Q10 2)
Q13 A)
Q 14 3)
Q15 C)

JEE-Advance

Q1. 9
Q2. 6

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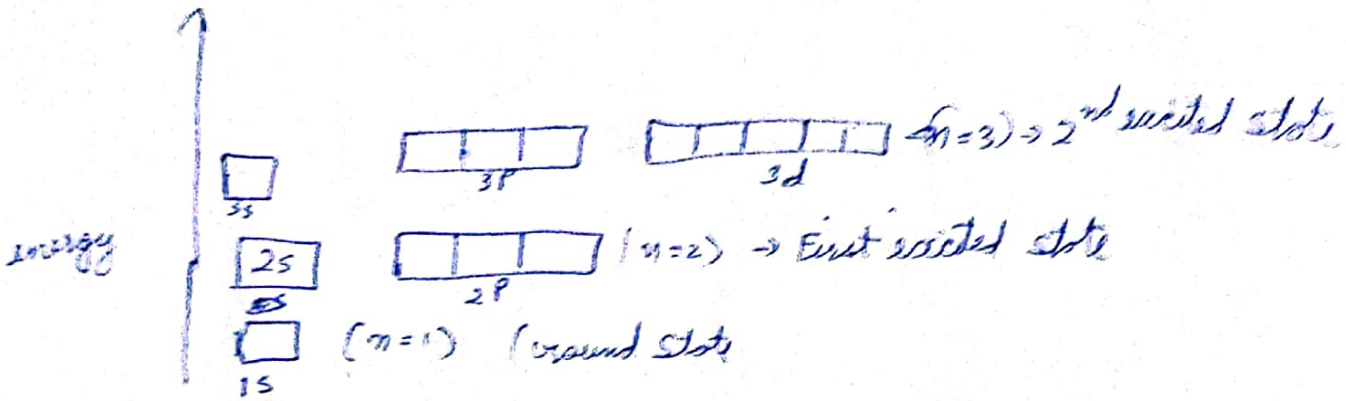
Single Electronic Species (H-like atom species)

- no. of $e^- = 1$

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

→ Energy will depend only on value of n

Energy Order- $1s < (2s = 2p) < (3s = 3p = 3d) < \ldots$



energy level diagram for single e^- species

eg. degeneracy of $n=2$ for H-atom $= 4$

degeneracy of 2nd excited state of H-atom $= 9$

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

degeneracy $= n^2$
$= (4)^2$
$= 16$

Multi-Electronic Species

no. of $e^- > 1$

eg. He^-, H_2, Li, Be ....

energy of e^- depends on value of n & l

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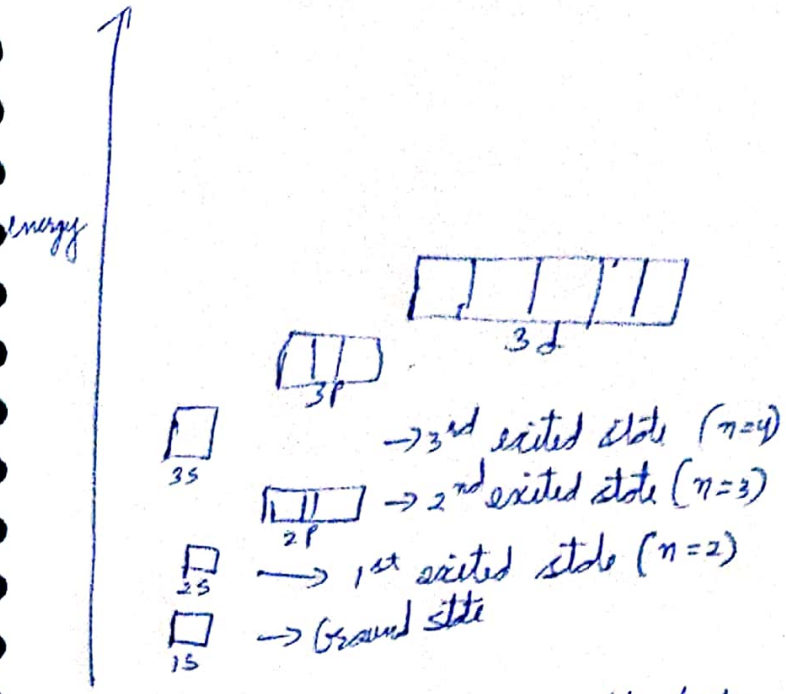

energy diagram for multi electronic species

Q what is the degeneracy of 2nd excited state of helium - 3

Q what is the degeneracy of 3rd excited state of Lithium - 1

Hunts 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 so S_z give maximum no. of e^- unpaired e^- with parallel spin

→ Pairing of e^- do not take place in orbitals until each orbital of the same subshell is filled with $1e^-$ of same spin.

Or

→ According to this rule, e^-s are filled into diff orbitals of a subshell so S_z give maximum spin multiplicity.

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1	1	1	1	1

✓

1	1	0	1	1

✓

1	↓	1	1	1

✗

↑	1	1	1	1

✓

↑	1	1	1	

✗

↑	1	1	1	↑

✓

Total Spin (S_T) :-

1	1	1

$|S_T| = |+\frac{1}{2}+\frac{1}{2}+\frac{1}{2}| = \frac{3}{2}$

↑	1	1

$|S_T| = |+\frac{1}{2}-\frac{1}{2}+\frac{1}{2}+\frac{1}{2}| = \frac{1}{1}$

Spin Multiplicity (S_m)

$$\boxed{S_m = 2|S_T|+1}$$

1	1	1

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

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

$\underline{S_m = 4}$

↑	1	1

$|S_T| = 1$

$S_m = 2\times1+1$

$\underline{S_m = 3}$

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Q1. The orbital diagram in which violated Afbau's principal is

i) [↑↓] [↑↓ | ↑ |]

ii) [↑] [↑↓ | ↑ | ↑] ✓

iii) [↑↓] [↑ | ↑ | ↑]

iv) [↑↓] [↑↓ | ↑ | ↓]

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

i) [↑ | ↑ | ↑ | | ↑]

ii) [↑ | ↑ | | ↓] ✓

iii) [↑↓ | ↑↑ |] ✓

iv) [↑ | ↑ | ↑ | ↑↓ | ↑]

Q3. Match

Configuration			Violated Rule
~~i) [↑↓ ↑↓]~~			
A i) [↑↓]	[↑↓]	[↑ | ↑]	P) Pauli's Rule
B ii) [↑↓]	[↑↑]	[↑ | ↓ |]	Q) No Rule violated
C iii) [↑]	[↑↓]	[↑ | ↓ | ↑]	R) Afbau's Rule
D iv) [↑↑]	[↑]	[↑ | ↓ | ↑]	S) Hund's Rule

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

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Electronic Configuration of a atom.

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

eg. $1s^2$ → no. of electrons

$n=1$; $l = s = 0$

$s \rightarrow 2e^-$
$p \rightarrow 6e^-$
$d \rightarrow 10e^-$
$f \rightarrow 14e^-$

eg.

$_1H \rightarrow 1s^1$

$_2He \rightarrow 1s^2$

$_3Li \rightarrow 1s^2 2s^1$

$_4Be \rightarrow 1s^2 2s^2$

$_5B \rightarrow 1s^2 2s^2 2p^1$

$_6C \rightarrow 1s^2 2s^2 2p^2$

$_7N \rightarrow 1s^2 2s^2 2p^3$

$_8O \rightarrow 1s^2 2s^2 2p^4$

$_9F \rightarrow 1s^2 2s^2 2p^5$

$_{10}Ne \rightarrow 1s^2 2s^2 2p^6$

Orbital notation method

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$_6C \rightarrow$ [1s] [2s] [2p]] orbital diagram method

$_6C \rightarrow [He]\ 2s^2\ 2p^2$] condensed form

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

$_{12}Mg \rightarrow [Ne]\ 3s^2$

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

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

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

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

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

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

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

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

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

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

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

$_{24}Cr \rightarrow$ ~~[Ar] 4s² 3d⁴~~ $[Ar]\ 4s^1\ 3d^5$

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

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

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

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

$_{29}Cu \rightarrow$ ~~[Ar] 4s² 3d⁹~~ $[Ar]\ 4s^1\ 3d^{10}$

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

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Inert Gas	Atomic no.
He	2
Ne	10
Ar	18
Kr	36
Xe	54
Rn	86
Uuo	118

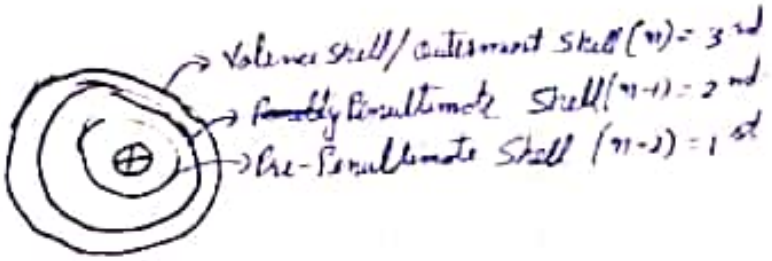
Exception

$_{24}Cr$:- [Ar] $4s^1$ $3d^5$

$_{29}Cu$ - [Ar] $4s^1$ $3d^{10}$

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.



Na - $1s^2 2s^2 2p^6 3s^1$

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Q1. write down the all four quantum No. for last e^- of Fe (iron)?

$_{26}Fe \rightarrow 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^6$

$\cancel{v\ e^-} \rightarrow n = 4$
$l = 0$
$m = 0$
$s = +\frac{1}{2} / -\frac{1}{2}$

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

Q2. Write write down the all four Q. No. for 7th e^- of Cl (atom)

$_{17}Cl \rightarrow 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$

$\cancel{7\ e^-} \rightarrow n = 3$
$l = 1$
$m = -1/-1/0$
$s = +\frac{1}{2} / -\frac{1}{2}$

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

Q3. find ~~minimum~~ maximum no. of e^- in sulfur for which $m = 0$?

$_{16}S \rightarrow 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$

for $m = 0$, $l > 0$ ~~any~~ d 2in every subshell.

$\boxed{10}$

Q4. find total no. of e^- in manganese for which

$_{25}Mn \rightarrow 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^5$

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

b) value of $l = 1$
$= 2 + 6$
$\boxed{= 12}$ ✓

c) value of $|m| = 1$
~~$= 2 + 2 + 2$~~
$= 4 + 4 + 2$
$\boxed{= 10}$ ✓

d)

e) value of $(n + l) = 4$
$= 6 + 2 +$
$\boxed{= 8}$ ✓

e) value of $(l + m) = 0$
$= 2 + 2 + 2 + 2 + 2 + 2 + 1$
~~$\boxed{= 13}$~~ $= \boxed{13}$ ✓

f) value of $(l + m) = 2$
$+ 2 + 2 + 1$
$\boxed{= 5}$ ✓

g) $(n + l + m) = 0$
$= 8 + 0 + 0$
$\boxed{= 13}$ ✓

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Q find out minimum no. of e^- in Magnesium for $s = -1/2$

$10+2 \rightarrow 12-$ $\boxed{4s}$

min no. of $e^- = 10$

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 those subshells which have more value of l (Azimuthal Quantum number)

Eg Na^+

$Na \rightarrow 1s^2\ 2s^2\ 2p^6\ 3s^1 \rightarrow$ outermost

$Na^+ \rightarrow 1s^2\ 2s^2\ 2p^6$

$Mn \rightarrow 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^5$

$Mn^{2+} \rightarrow 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^5$

$Mn^{4+} \rightarrow 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^3$

$Mn^{7+} \rightarrow 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

$N \rightarrow 1s^2\ 2s^2\ 2p^3$

$N^{3-} \rightarrow 1s^2\ 2s^2\ 2p^6$

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Extra stability of half filled & fully filled subshell.

1. The electronic configuration of most of the elements follows Aufbau's rule.

Exception of Aufbau's Rule:-

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

Ar [↑↓] [↑|↑|↑|↑|↑] → [↑↓] [↑|↑|↑|↑|↑]

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

Ar [↑↓] [↑↓|↑↓|↑↓|↑↓|↑↓] → [↑] [↑↓|↑↓|↑↓|↑↓|↑↓]

2. Here the two subshell (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 result in all orbitals of a subshell of higher energy getting either completely filled or half filled

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Half filled subshell	Fully filled subshell
s - [↑]	[↑↓]
p - [↑][↑][↑]	[↑↓][↑↓][↑↓]
d = [↑][↑][↑][↑][↑]	[↑↓][↑↓][↑↓][↑↓][↑↓]
f - [↑][↑][↑][↑][↑][↑][↑]	[↑↓][↑↓][↑↓][↑↓][↑↓][↑↓][↑↓]

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 stabilisation is due to following 2 factors

Explanation 1- Symmetry leads to stability (Nature's law)

Exp 2 - <u>Exchange energy</u> - 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.

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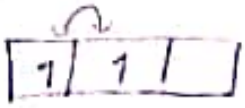
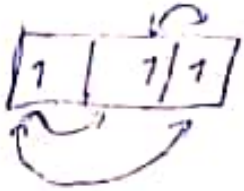


no of exchange = 1



no. of exchange = 3

no. of exchange = 4

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

n = no of e^- in up spin

n' = no. of e^- in down spin

No of exchange ↑ ⇒ energy ↓ ⇒ stability ↑

---

* <u>Excited energy</u> – 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.

eg.

C – [↑↓] [↑↓] [↑|↑|] → Ground State ($1s^2\ 2s^2\ 2p^2$)

C* : [↑↓] [↑] [↑|↑|↑] → excited state ($1s^2\ 2s^1\ 2p^3$)

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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	$[Kr]\ 5s^1\ 4d^4$
4	Mo	42	$[Kr]\ 5s^1\ 4d^5$
5.	Ru	44	$[Kr]\ 5s^1\ 4d^7$
6.	Rh	45	$[Kr]\ 5s^1\ 4d^8$
7	Pd	46	$[Kr]\ 4d^{10}$
8	Ag	47	$[Kr]\ 5s^1\ 4d^{10}$
9.	La	57	$[Xe]\ 6s^2\ 5d^1$
10	Pt	78	$[Xe]\ 6s^1\ 4f^{14}\ 5d^9$
11.	Au	79	$[Xe]\ 6s^1\ 4f^{14}\ 5d^{10}$
12	Ac	89	$[Rn]\ 7s^2\ 6d^1$
13.	Th	90	$[Rn]\ 7s^2\ 6d^2$

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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 (Q1, 2, 3, 5, 6)

S-2 (Q1, 2, 3, 9)

JM (Q2, 4, 12, 16)

JA (Q3)

JA Q3 - 3

JM Q2 - 1)
Q4 - 3)
Q 12 - 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 (Doubt)
Q6 - 2 (no. of orbitals)

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 - C
Q 16 - D
Q 17 - A

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Q Find maximum no. of e^- in 5th period.

$n = 5$

5s , (5-3) g , (5-2) f , (5-1) d , 5p

5s , ~~g~~ , ~~f~~ , 4d , 5p

2 , 10 , 6

10 + 6 + 2 → 15 elements

Magnetic Properties

Paramagnetic

Diamagnetic

Spin Magnetic moment (μ) -

μ (mew)

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

→ $\boxed{\mu = \sqrt{(n)(n+2)}}$

n: no. of unpaired electrons

→ unit - B.M. (Bohr Magneton)

$_{12}Mg$ - $1s^2\ 2s^2\ 2p^6\ 3s^2$

[↑↓] [↑↓] [↑↓ ↑↓ ↑↓] [↑↓]

no of unpaired electron = n = 0

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Eg2- $_7N \rightarrow 1s^2\ 2s^2\ 2p^3$

[↑↓] [↑↓] [↑|↑|↑]

$n = 3$

$\mu = \sqrt{3(3+2)}$

$\mu = \sqrt{15}$

$\mu = 3.87$

no. of unpaired e^-	$\sqrt{n(n+2)}$	μ
0	0	0
1	$\sqrt{3}$	1.73
2	$\sqrt{8} = 2\sqrt{2}$	2.84
3	$\sqrt{15}$	3.87
4	$\sqrt{24}$	4.91

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

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

$4s^2\ 3d^7$

[↑↓] [↑↓|↑↓|↑|↑|↑]

+1 → [↑] [↑↓|↑↓|↑|↑|↑] → $n = 4$ $x = 1$

+2 → [ ] [↑↓|↑↓|↑|↑|↑] → $n = 3$ $x = 2$

+6 → [↑|↑|↑| |] → $n = 3$ $x = 6$

$\boxed{x = 2, 6}$

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Magnetic

Paramagnetic	Diamagnetic
1 weakly attracted by external magnetic field.	1. weakly repelled by external magnetic field.
2. $\mu \neq 0$	2. $\mu = 0$
3. At least 1 unpaired e^- will be present	3. no unpaired e^- present
eg. N, Cl, C	eg. Ne, Mg, ~~Ca~~, Na^+, F^-, N^{3-}

Shape of orbitals & their orientation:-

1. Nodal Plane - It is defined as the plane passing through the nucleus where probablity 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

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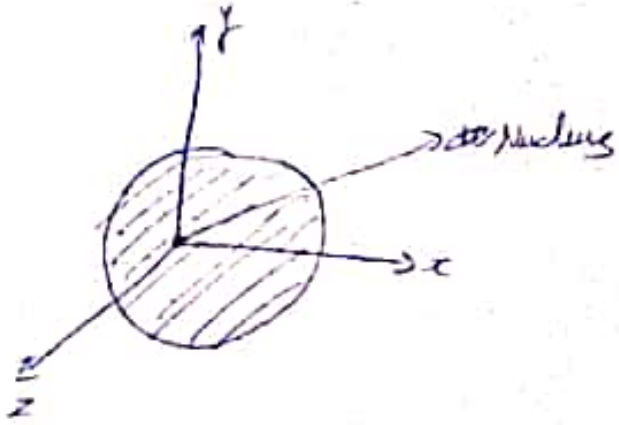
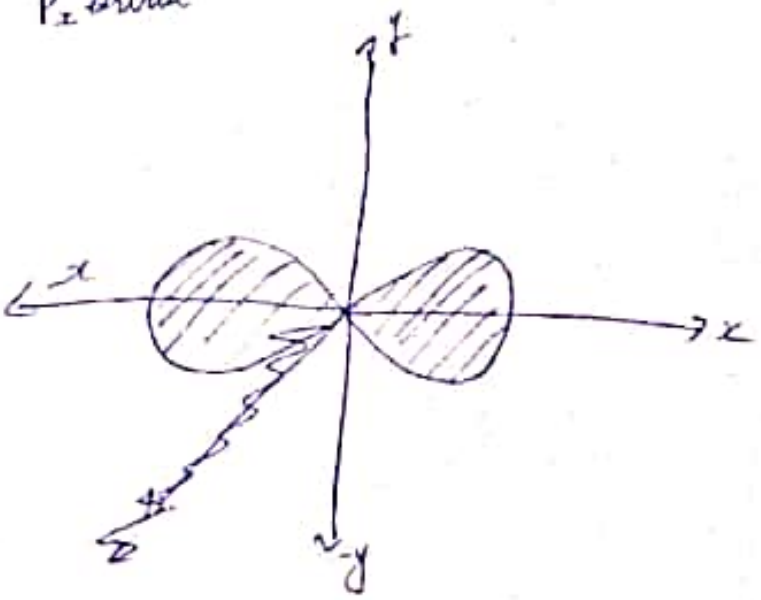
1. S-orbital (spherical)



N.P = 0

2. P-orbital (Dumbell)

P_x orbital



N.P = 1

Nodal plane - YZ axis or ZY

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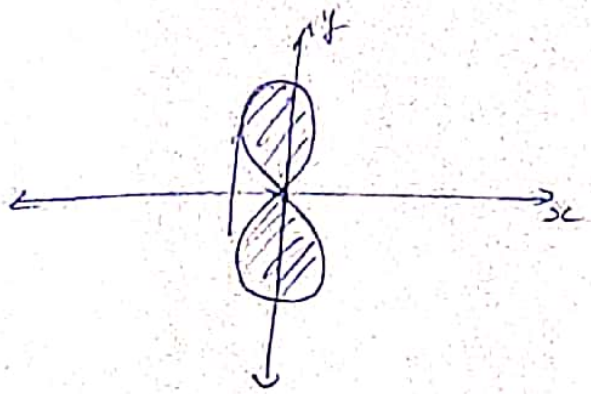
P_y orbital -



Nodal Plane - XZ or ZX

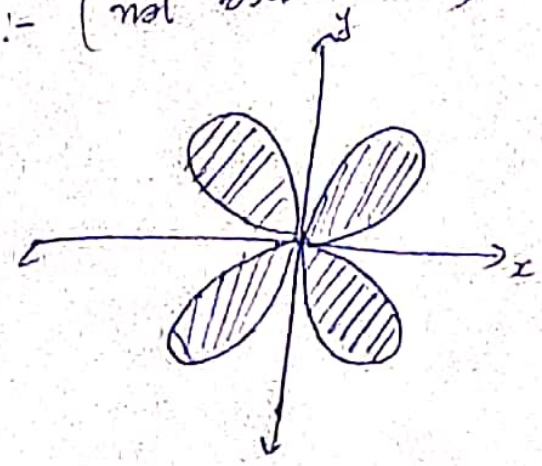
These orbitals are axial orbitals because e^- density lies of axis.

P_z orbital:-

Nodal plane - XY or YX

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

d_{xy} :- (not axial orbital)



Nodal Plane - XZ and YZ

-2

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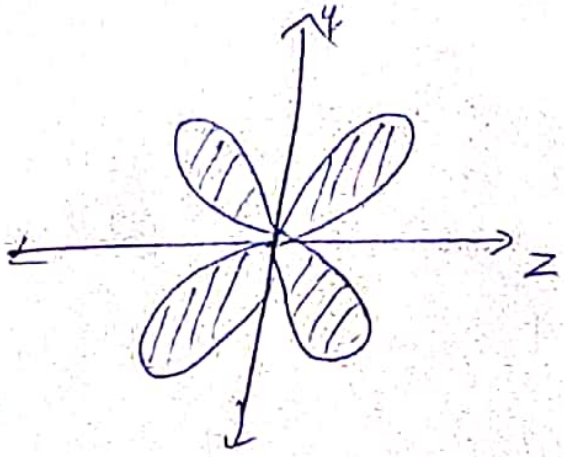
d_{yz} (not axial orbital



Nodal Planes :- XY, XZ

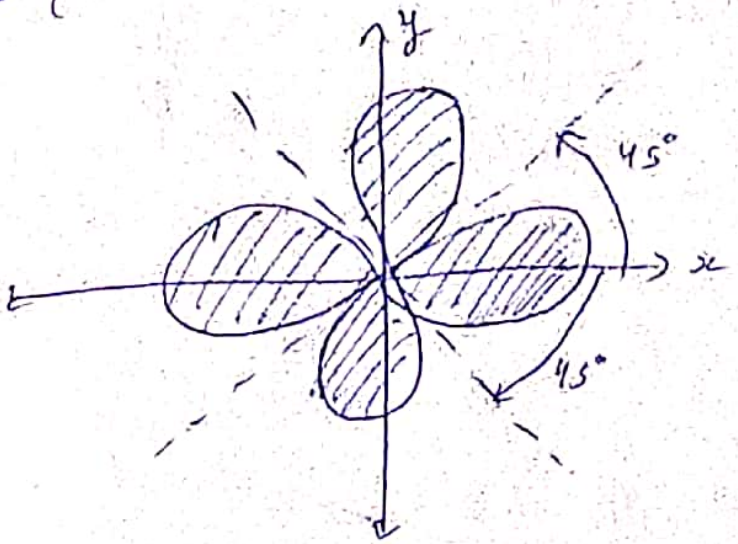
- 2

d_{xz} :- (not axial orbital)

Nodal Planes :- YZ, XY

- 2

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



Nodal Planes :- 2
- Two nodal plane inclined at 45° from both X & Y axis

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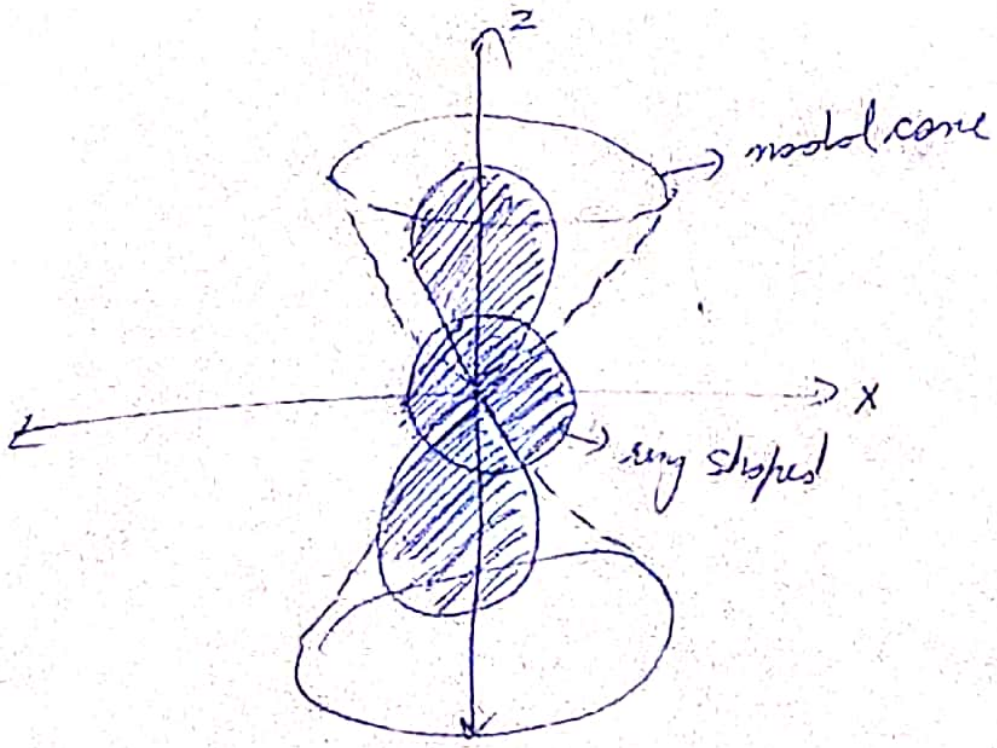
☆ d_{z^2} :- (Axial orbital)



Nodal Planes :- 0
- No nodal planes
- two nodal cones exist

Screening constant & effective nuclear charge (Z^*/Zeff)
(σ)
↓
Sigma

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

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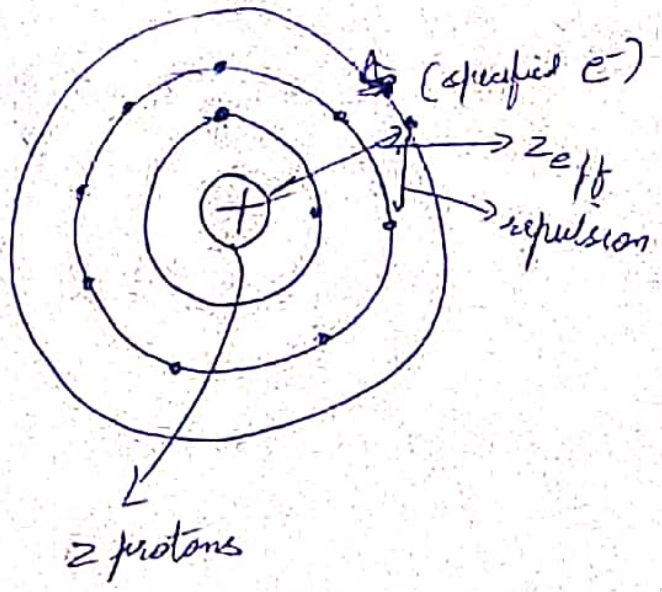
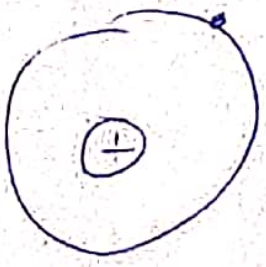


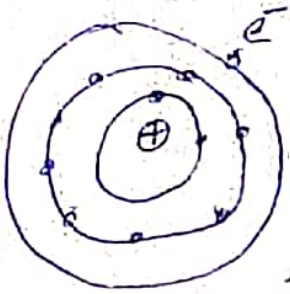
for single e⁻ system



ē specified ⇓ force of attraction on specified e⁻ of Z no. of protons

$Z_{eff} = Z$

for multi e⁻ system



e⁻ ⇒ force of attraction on specified ē of Z no. of protons

$Z_{eff} = Z - \sigma$ → Repulsion

Z_{eff} → net attraction force; Z → Total attraction

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Z_{eff} = effective Nuclear charge

Z = atomic no. / no. of proton

σ = slater constant / shielding constant / screening constant

---

Note:-

→ Due to the presence of inner electron, 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 confign of an atom and arrange ec 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.

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2 Types of groups

ns/np

1. All the e^- present right hand side of the specified e^- contribute nothing to sigma.
2. The remaining e^- in same group (ns/np) 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 1 each to sigma.

nd/nf

1. e^- present right hand side of Nd or Nf grp contribute nothing to sigma.
2. e^- present in same group (nd/nf) contribute 0.35 each to sigma.
3. All e^- present in LHS of (nd/nf) grp contribute 1 each to σ.

Zeff for (4s) 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 Zeff on 3d is more.

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Q1. Calculate value of σ & Zeff for valence shell e^- of Be.

Be - $(1s^2)^2(2s^2)$ (2s) → specified e^- / test e^-

$\sigma = 0.35 + 2(0.85)$

$\sigma = 0.35 + 1.7$

$\sigma = \boxed{2.05}$

$Z_{eff} = Z - \sigma$

$Z_{eff} = 4 - 2.05$

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

~~Q2. Calc...~~

Q2. Calculate value of Zeff & σ for 3s e^- of Na

Na → $(1s)^2(2s^2 2p^6)$ (3s¹)

$\sigma = (8 \times 0.85) + (1 \times 2) + (0 \times 0.35)$

$\sigma = 6.80 + 2$

$\sigma = 8.8$

$Z_{eff} = Z - \sigma$

$= 11 - 8.8$

$\boxed{= 2.2}$

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Q3. Calculate value of Z_{eff} & σ for 3d e^- of Iron. also calculate σ & Z_{eff} for 4s e^-

$Fe - (1s)^2 \ (2s, 2p)^8 \ (3s, 3p)^{8} \ (3d)^6 \ (4s)^2$

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

$= 18 + 1.75$

$\boxed{= 19.75}$

$3d \ Z_{eff} = Z - 19.75$

$= 26 - 19.75$

$\boxed{= 6.25}$

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

~~$= 18 + 0.35 + 5.10$~~

~~$\boxed{= 23.45}$~~

~~$4s \ Z_{eff} = 26 - 23.45$~~

~~$\boxed{= 2.65}$~~

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

$= 10 + 0.35 + 11.9$

$\boxed{= 22.25}$

$4s \ Z_{eff} = 26 - 22.25$

$\boxed{= 3.75}$

$$\begin{array}{r} 2 \\ 85 \\ 14 \\ \hline 340 \\ 850 \\ \hline 11.90 \end{array}$$

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$_3Li$ - $(1S)^2$ $(2S)^1$

$\sigma = 0.85 \times 2$

$= 1.7$

$Z_{eff} = Z - \sigma$

$= 3 - 1.7$

$\boxed{= 1.3}$

$_4Be$ - $(1S)^2$ $(2S)^2$

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

$= 4 - 2.05$

$\boxed{= 1.95}$

$_5B$ - $(1S)^2$ $(2S, 2P)^3$

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

$= 5 - 2.4$

$\boxed{= 2.6}$

$_6C$ - $(1S)^2 (2S, 2P)^4$

$Z_{eff} = 6 - (2 \times 0.85 + 3 \times 0.35)$

$= 6 - 2.75$

$\boxed{= 3.25}$

$_7N$ - $(1S)^2$ $(2S, 2P)^5$

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

$= 7 - 3.1$

$\boxed{= 3.9}$

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$_8O$ - $(1S)^2$ $(2S, 2P)^6$

$Z_{eff} = 8 - (2 \times 0.85 + 5 \times 0.35)$

~~$= 8 - (1.7 + 2.10)$~~

~~$= 8 - 3.8$~~

$= 8 - (1.7 + 1.75)$

$= 8 - 3.45$

$\boxed{= 4.55}$

$_9F$ - $(1S)^2$ $(2S, 2P)^7$

$Z_{eff} = 9 - (2 \times 0.85 + 6 \times 0.35)$

$= 9 - 3.8$

$\boxed{= 5.2}$

$_{10}Ne$ - $(1S)^2$ $(2S, 2P)^8$

$Z_{eff} = 10 - (2 \times 0.85 + 7 \times 0.35)$

$= 10 - 4.15$

$\boxed{= 5.85}$

$_{11}Na$ - $(1S)^2$ $(2S, 2P)^8$ $(3S)^1$

$Z_{eff} = 11 - (2 \times 1 + 8 \times 0.85 + 0 \times 0.35)$

$= 11 - (2 + 6.80 + 0)$

$= 11 - 8.8$

$\boxed{= 2.2}$

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Q Calculate the Zeff of 3d e^- of Cr^{3+} ?

$_{24}Cr^{3+}$ $1s^2$ $2s^2$ $2p^6$ $3s^2$ $3p^6$ ~~4s^0~~ $5d^3$

$(1s)^2 (2s, 2p)^8$ $(3s, 3p)^8$ $(5d)^3$

$\sigma = (7 \times 0.35) + (8 \times 0.85) + 2$

$= 2.45 + 3.60 + 2$

$= 6.05 + 2$

$= 18.05$

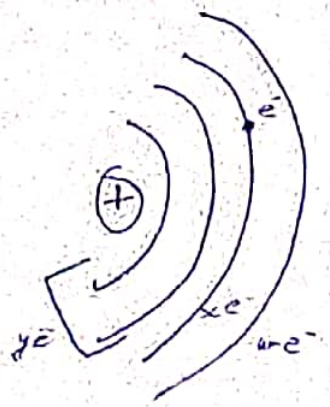
$Z_{eff} = \cancel{24 - 5.05}$ $24 - 18.7$

$\cancel{= 18.95}$ $\boxed{5.3}$

for ns, np e^-

$\sigma = (w \times 0) + (x-1) \times 0.35 + (y \times 0.85) + z \times 1$

for nd or nf e^-



$\sigma = (w \times 0)(x-1) \times 0.35 + y \times 1$

Note:- for $1s^2$ e^- configuration :-

$\sigma = 0.3$

Eg. He - $1s^2$

$\sigma = 0.3$

$Z_{eff} = 2 - 0.3$

$= 1.7$

Li^+ - $1s^2$

$\sigma = 0.3$

$Z_{eff} = 3 - 0.3$

$= 2.7$

Ni - $1s^2 \underbrace{2s^2 2p^6 3s^2 3p^6 4s^2 3d^8}_{x}$

$\sigma = 28.3$

$Z_{eff} = 28 - 0.3$

$= 27.7$

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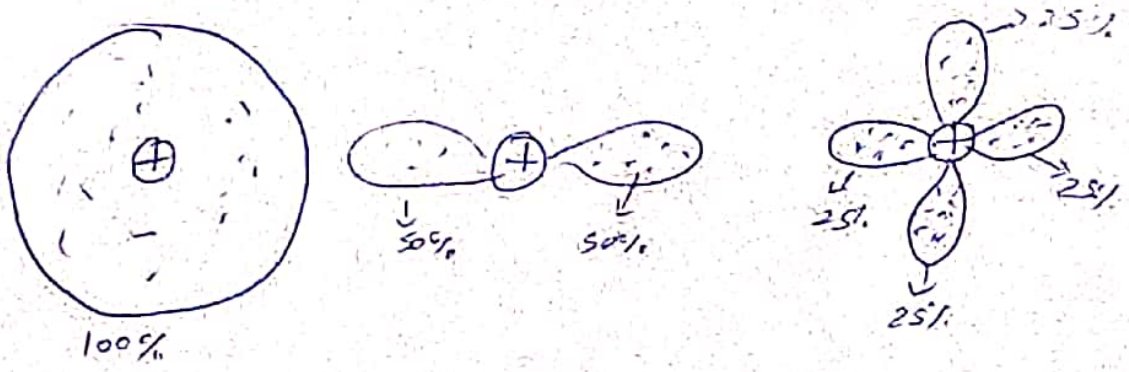
Note :- For H atom, $Z_{eff} = 1$

<u>Penetration Power</u> - It is the ability of an orbital to attract e^-

order - $ns > np > nd > nf$

(How easy can electron approach to nucleus)



<u>Sheilding Power / Screening power</u> :-

order - $ns > np > nd > nf$

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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. Dobereiner's – Triad rule
3. Newlands – Rule of octave
4. Lothar – Mayer curve
5. Mendeleev's – Periodic table
6. Mossely's – Experiment off and modern periodic table.

1. **Lavoisier Classification** – firstly he classified elements into 2 categories

Element
- Metal
- Non-metals

2. **Dobereiner's Triad rule** – He made the group of 3 elements having similar chemical properties called ~~Triad~~ triad.

→ In Dobereiner's triad, atomic mass of central element is nearly equal to average atomic mass of 1st & 3rd element

Li	Na	K
Be	Mg	Ca
Cl	Br	I

Li	Na	Mg
7	23	39

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

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K Rb Cs

S Se Te

Ca Sr Ba

~~F Cl Br~~ → not following

Cl Br I

35.5 81.25 127

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

F Cl Br

19 35.5 80

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

Q1. Multi Correct

If three element ^{20}X ^{42}Y & ^{0}A Z follow Doberiner'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	Re	Ga	Ma	Pa	Dha	Ni	Sa

Li Be B C N O F

Na Mg Al Si P S Cl

K Ca Sc

☆☆ Nobel gases were not yet discovered.

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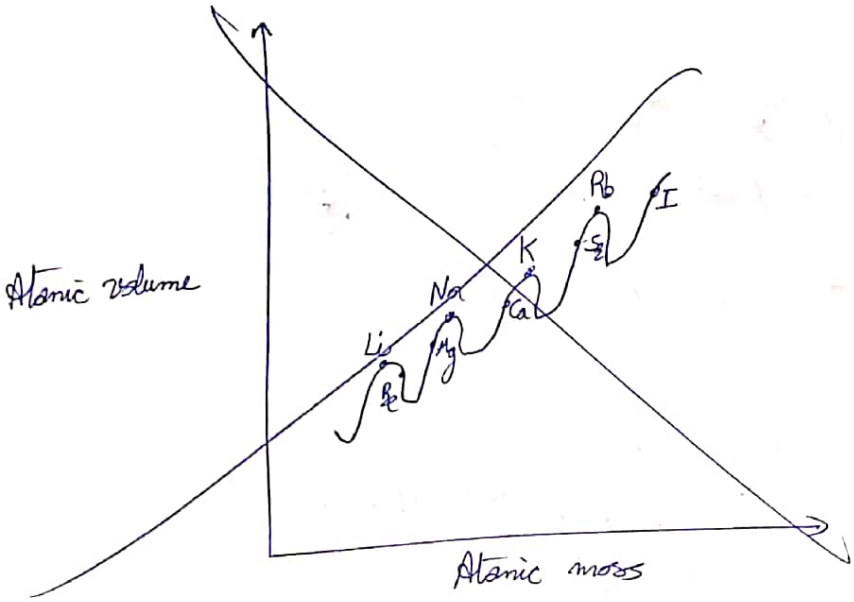
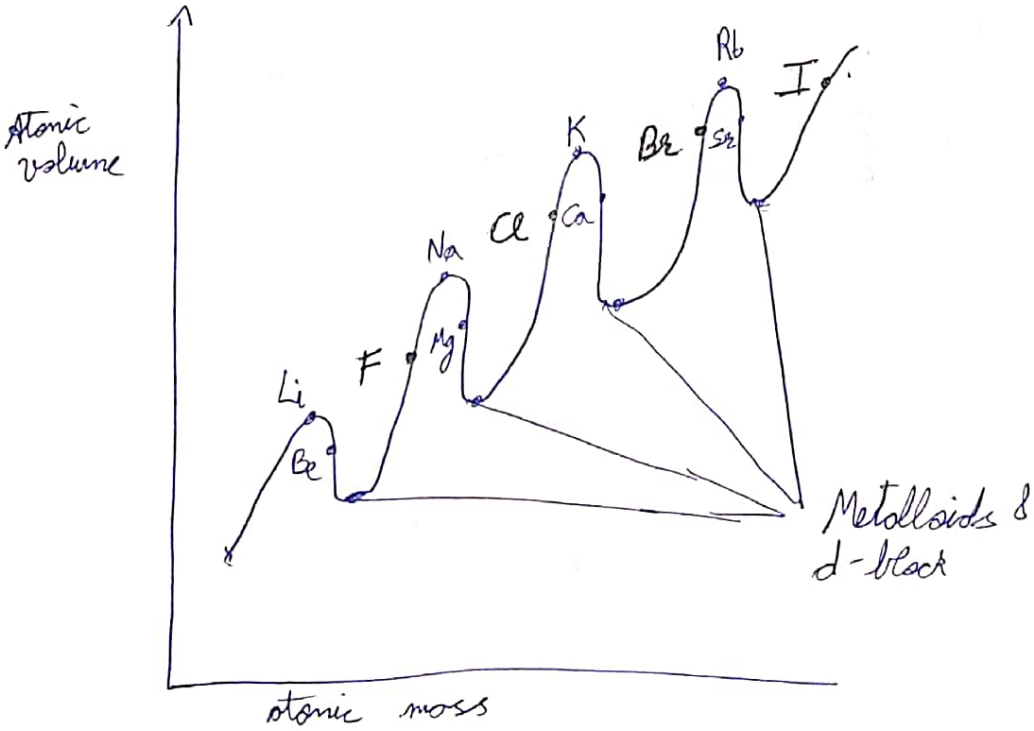
merits:- He classified the known elements in tabular form.

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

:- discovery of noble gases failed the rule.

4. Lothar Meyer's curve -





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Observations 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 decending portion of the curve.

→ Halogens (F, Cl, Br, I) occupied ascending portion of the curve.

→ Metaloids & transition elements occupied lower portions of the curve

Drawbacks → It is highly complex when further 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.

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P \ G	I		II		III		IV		V		VI		VII		VIII		
	A	B	A	B	A	B	A	B	A	B	A	B	A	B			
1	H																
2	Li		Be		B		C		N		O		F				
3	Na		Mg		Al		Si		P		S		Cl				
4	K	Cu	Ca	Zn	Ga	Sc	Ge	Ti	As	V	Se	Cr	Br	Mn	Fe	Co	Ni
5	Rb	Ag	Sr	Cd	In	Y	Sn	Zr	Sb	Nb	Te	Mo	I	Tc	Ru	Rh	Pd
6	Cs	Au	Ba	Hg	[illegible]		Pb		Bi		Po		[illegible]		Os	Ir	Pt
7	Fr		Ra														

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

1. Around 63 elements were arranged in order of increasing atomic weight into 7 horizantal rows (called periods) and 8 vertical columns (called groups).
2. Each group except 8th group was further divided into 2 sub-groups
3. Subgroup A 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 same group have similar properties.

Merits - 1. ~~old known~~ all known elements were arranged.
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 - Mangenese (Mn)

Modern Name

Gallium (Ga)
Germanium (Ge)
Scandium (Sc)
Technetium (Tc)

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Si - Ge और Al - Ga
गीता - गीता अलग

B - Sc Mn - Tc
B.Sc नहीं M. Tech है।

3. Correction of doubtfull atomic weights

U, Be, In, Au, Pt
यू, बेईमान, इंसान, Gold और Platinum पुराने हैं।

Demerits-

1. Position of Hydrogen(H) was not fixed properly
2. Position of Isotopes was not confirmed
3. There were certain pairs which did not obey their increasing weight order, such pairs are called anomalous pairs.

~~eg. (Ar ⟷ K) 39.9 39.1~~
~~(Te ⟷ I) 127.5 127~~
~~Th ⟷ Pa 232 231~~
~~Co ⟷ Ni 58.9 58.6~~

eg. आमिर खान
Ar ⟷ K
39.9 39.1

मेरी इच्छा
Te ⟷ I
127.5 127

ठंडी पेप्सी
Th ⟷ Pa
232 231

Coke नहीं
Co ⟷ Ni
58.9 58.6

4. Like elements were placed in different groups
eg. (Au – Pt)
Au → IB, Pt → VIII

5. Unlike elements were placed in same group.

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eg.

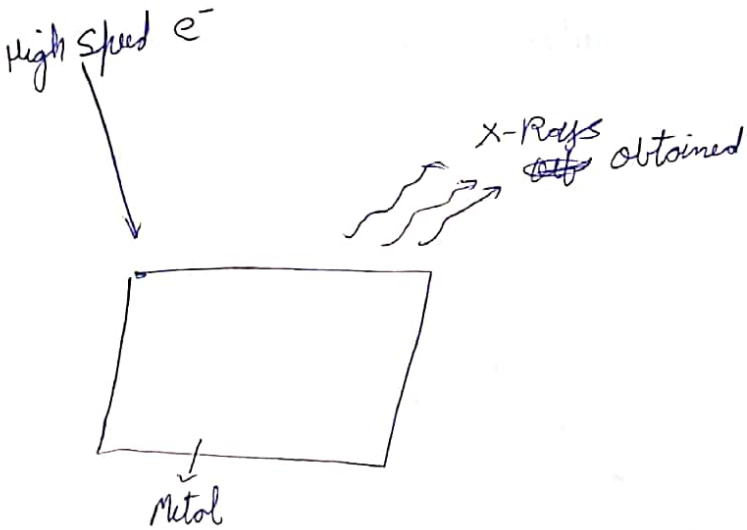
I-A

Li, Na, K, Rb, Cs — Highly reactive

IB

Cu, Ag, Au — coinage metals (less reactive)

∴ 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.

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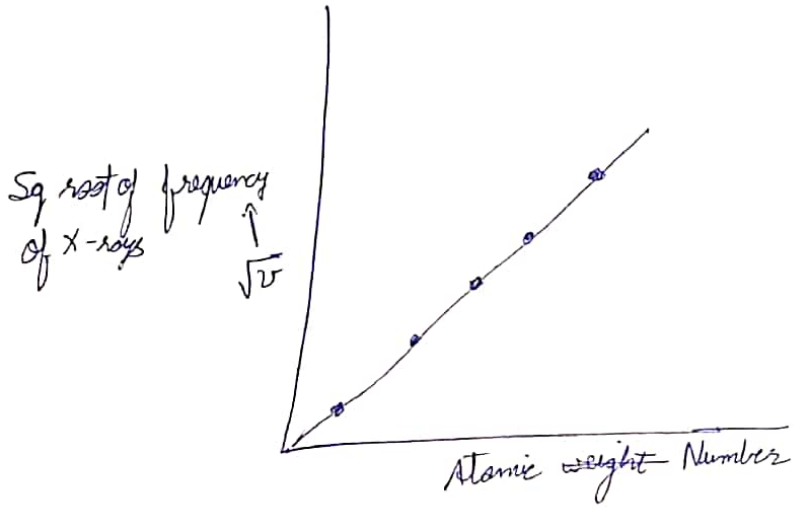

$\sqrt{\nu} \propto Z$

$$\boxed{\sqrt{\nu} = a(Z-b)}$$

Z - atomic number

ν = frequency of X-ray

a, b = moseley's constant

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

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

$\sqrt{625} = 25$

$25 = a(z-b)$

$25 = 1(z-2)$

$25 = z-2$

~~z = 27~~

$z = 27$

$\boxed{\text{Cobalt (Co)}}$

Modern Periodic Table

1. 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.
2. It consists of total 18 vertical columns called Groups, and 7 horizontal rows called periods.

Deciding the block of an element -

1. There are 4 blocks in a periodic table which are S, P, D & F.
2. To decide the block we first write the electronic configuration of the element.
3. The type of subshell in which last electron enters will give the information about block of that element. (Exception - $_{90}Th$ (Thorium)

~~$1s^2 2s^2$~~

$[Rn]_{86} 6d^2 7s^2$

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

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eg. $_{11}Na$ - $1s^2 2s^2 2p^6 3s^1$

$\boxed{\text{S- Block}}$

$_{24}Cr$ - $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^5$

$\boxed{\text{D- block}}$

$_{7}N$ - $1s^2 2s^2 2p^3$

$\boxed{\text{P-Block}}$

Group I (IA)

Config:- ns^1 (valence shell $e^- = 1$)

$1s^1$ H]— Non-metal

है

$2s^1$ Li
$3s^1$ Na
$4s^1$ K
$5s^1$ Rb
$6s^1$ Cs
$7s^1$ Fr

— Alkali metals (Highly reactive)

ली
ना
की
रब
से
फरियाद

Group II (IIA)

Config ns^2 (valence shell $e^- = 2$)

$2s^2$ Be
$3s^2$ Mg
$4s^2$ Ca
$5s^2$ Sr
$6s^2$ Ba
$7s^2$ Ra

Alkaline earth methods

बेटा
मांगे
कार
स्कूटर
बाप
राजी

Scanned with CamScanner

Note - Group I & Group II are s-Block elements

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

Config - $ns^2 np^1$

$2p^1$	B		बैंगन
$3p^1$	Al		आलू
$4p^1$	Ga	→ Boron family	गाजर
$5p^1$	In		In
$6p^1$	Tl		~~थैला~~ थेला

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

Config - $ns^2 np^2$

$2p^2$	C		कट्टे	कट्टे
$3p^2$	Si	Carbon	गीता	शिव
$4p^2$	Ge	~~Fosbon~~	जी	जी
$5p^2$	Sn	Family	सुनो	सुनो
$6p^2$	Pb		प्र मु	पारकी

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

Config - $ns^2 np^3$

$2p^3$	N	Nitrogen family	Nepal
$3p^3$	P	(pnicogens)	Pakistan
$4p^3$	As		Australia
$5p^3$	Sb		Sab
$6p^3$	Bi		Bikhari

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

Config - $ns^2 np^4$

$2p^4$	O		old
$3p^4$	S	oxygen	style
$4p^4$	Se	family	से
$5p^4$	Te		ती
$6p^4$	Po	(chalcogens)	पो

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

Config - $ns^2\ np^5$

$2p^5$	F	Halogens family	फ्लोर
$3p^5$	Cl		क्लोर
$4p^5$	Br		ब्रोमिन
$5p^5$	I		आयोडिन
$6p^5$	At		Astaty / Astatine

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

Config - $ns^2\ np^6$

$1s^2$	He (valence shell $e^- = 2$) (s-block element)	Noble gases (inert gases)	हीलियम
$2p^6$	Ne		नीयॉन
$3p^6$	Ar		आर्गन
$4p^6$	Kr		क्रिप्टॉन
$5p^6$	Xe		जेनॉन
$6p^6$	Rn		रेडॉन

Note:-

1. Group 13, 14, 15, 16, 17 & 18 are known as P-Block elements except (Helium)
2. Helium is S-Block element.
3. 18th group is known as noble gases/inert gases.
4. Except group 18th 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.

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D- block elements :-

① 3d-series or I^{st} transition series elements.

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

Science Teacher Vijay Cryptoni मान के कोनी कुछ जान जलाते हैं।

② 4d-series or 2nd ~~Transition~~ Transition series elements.

$_{39}Y$ $_{40}Zr$ $_{41}Nb$ $_{42}Mo$ $_{43}Tc$ $_{44}Ru$ $_{45}Rh$ $_{46}Pd$ $_{47}Ag$ $_{48}Cd$

यारी जरा निमा मौत तक रूकावट राह पड़े आगा में कूद जा।

③ 5d-series or 3rd Transition elements.

$_{57}La$ $_{72}Hf$ $_{73}Ta$ $_{74}W$ $_{75}Re$ $_{76}Os$ $_{77}Ir$ $_{78}Pt$ $_{79}Au$ $_{80}Hg$

ला हफता वरना रे ओसामा ~~Editor~~ इधर पिटाई और होगी

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.

Zn

$Zn : [Ar]\ 4s^2\ 3d^{10}$

$Zn^{2+} : [Ar]\ 3d^{10}$

not Transition element

$Cu : [Ar]\ 4s^1\ 3d^{10}$

$Cu^{+} : [Ar]\ 4s^0\ 3d^{10}$

$Cu^{2+} : [Ar]\ 4s^0\ 3d^9$ → Transition element

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F-Block elements

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

Lanthanoids / Rare earth metals

D-block element

$_{57}La$ ($_{58}Ce$ $_{59}Pr$ $_{60}Nd$ $_{61}Pm$ $_{62}Sm$ $_{63}Eu$ $_{64}Gd$

मीने पर नदिया प्रेम की सगाई यू गद गद

$_{65}Tb$ $_{66}Dy$ $_{67}Ho$ $_{68}Er$ $_{69}Tm$ $_{70}Yb$ $_{71}Lu$)

तब Dil Hua और तुम यूही लाजवाब हो।

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

$_{89}Ac$ ($_{90}Th$ $_{91}Pa$ $_{92}U$ $_{93}Np$ $_{94}Pu$ $_{95}Am$ $_{96}Cm$ $_{97}Bk$ $_{98}Cf$ $_{99}Es$

D-block element तांगे पर अमी बरस पड़ी आसमा चमका बिजली कड़की एहसान

$_{100}Fm$ $_{101}Md$ $_{102}No$ $_{103}Lr$)

फरामोश मनचला निकला लड़का

Note :-

1. Lanthanium & Actinium (La & Ac) are D- block elements
2. Lanthanite series elements & Actinite series elements are F-block elements.
3. Lanthanite series starts from $_{57}Ce$ to $_{71}Lu$
4. Actinite series starts from $_{90}Th$ to $_{103}Lr$

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

Eg. $1s^2\ 2s^2\ 2p^6\ 3s^2 \rightarrow Mg$

$\downarrow$

$n=3$

3rd Period

$Cr - 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1\ 3d^5$

$\downarrow$

4th period

No. of elements in a specific period.

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

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

ns .... (n-3)g (n-2)f (n-1)d np ~~for~~ 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 - $\boxed{\uparrow\downarrow\uparrow}$ - $3e^-$ → 3 elements

2P - $9e^-$ → 9 elements

Total elements → $\boxed{\text{12 elements}}$

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

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

6s → $3e^-$

4f → 14+7 → $21e^-$

5d → 10+5 → $15e^-$

6p → 6+3 → $9e^-$

Total e^- = 48

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

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Q3. If each orbital occupied by max $3e^-$, find max elements in 10th period.

$10s$	$6h$	$7g$	$8f$	$9d$	$10p$
$3e^-$	$33e^-$	$27e^-$	$21e^-$	$15e^-$	$9e^-$

$\rightarrow 108\ e^-$

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

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 = 135 th element~~ ~~+18~~

~~= 153 element~~

18th group → 50th element of period

17th group → 49th element of period

$118 + 49$ ~~= 168~~ $= 167$

$\boxed{Z = 167}$

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Q4. In other universe if all quantum nos. are same except l, $l = 0$ to n

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

ii) elements in 3rd shell

~~in 2nd~~

i) ~~In 2nd period → ns np~~

~~↓ ↓~~

~~$1e^-$ $1e^-$~~

~~→ 3 elements~~

~~3rd period → 3 elements~~

~~4th period → ns (n-1)d np~~

~~→ 3 e^-~~

~~→ 3 elements~~ 4 elements

ii) ~~3 elements~~ 4 elements

i) $n = 1$ 1s 1p

$n = 2$ 2s 2p 3d

$n = 3$ 3s 3p 3d 3f

2nd period :- 2s 2p

Total $e^- = 8$

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3rd period → 3S 2d 3P

→ 18 elements

4th period → ~~3S 4S~~ 4S 3d 4P

→ 18 elements

ii)

$n = 3$ → 3S 3P 3d 3f

2 6 10 14

32 elements

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

$l \rightarrow 0$ to $(n+1)$

$m \rightarrow -(l+1)$ to $+(l+1)$

$m_s \rightarrow -1/2, 0, +1/2$

Then, calculate

no. of elements in 2nd period

no. of elements in 2nd shell

2nd period → 2S 1d 2P

2 orbital 7 5

$3 \times 3 = 9e^-$ $3 \times 7 = 21e^-$ $3 \times 5 = 15e^-$ = $45e^-$

= 45 elements

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2nd Shell → $n = 2$

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

$m = -1, 0, 1 \rightarrow 3$
$-2, -1, 0, 1, 2 \rightarrow 5$
$-3, -2, -1, 0, 1, 2, 3 \rightarrow 7$
$-4, -3, -2, -1, 0, 1, 2, 3, 4 \rightarrow 9$

$= 24$ orbitals

$24 \times 3 \rightarrow 72 e^-$

$\boxed{\rightarrow 72 \text{ elements}}$

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Some Commonly used terms :-

1. Post Lanthanides elements – After Lanthanide elements (after Lu, Z = 71) or Hafnium (Z = 72) onwards elements are known as post Lanthanide elements.

   including

   not included

2. Trans Uranic elements – After Uranium element, atomic no. 92 (Z = 92) or (Np, Z = 93) are known as trans uranic elements.

   excluded

   including

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-Subshell,

   General electronic Config ns^{1-2}

   ns^1 or ns^2

   eg. Li – $1s^2 2s^1$

   Mg : [Ne] $3s^2$

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2. <u>P- Block elements</u> - last electron in P-subshell

E.C → ns^2 np^{1-6}

eg. N → $1s^2 2s^2 2p^3$

3. <u>D-Block elements</u> - last electron in D-subshell

E.C → ns^{0-2} $(n-1)d^{1 \text{ to } 10}$

eg. Cr → [Ar] $4s^1$ $3d^5$

Fe → [Ar] $4s^2$ $3d^6$

Pd → [Kr] $5s^0$ $4d^{10}$

4- <u>F-block elements</u> - last electron in F-subshell

E.C. → ns^{0-2} $(n-2)f^{1 \text{ to } 14}$ $(n-1)d^{0 \text{ to } 2}$

eg. $\underset{71}{Lu}$ - $\underset{54}{[Xe]}$ $6s^2$ $4f^{14}$ $5d^1$

<u>Trick to estimate group number of an electron</u>

→ To decide the group of ~~an~~ an element, first decide the block of the element.

→ <u>S-Block</u> - ns^1 → 1st group H, Li, Na, K, Rb, Cs, Fr

(Except Helium) ns^2 → 2nd group Be, Mg, Ca, Sr, Ba, Ra

no. of electron in valence shell = Group no. (for S-Block)

→ <u>P-Block</u> -

Group no. → 12 + np^{e^-}

eg S → [NE] $3s^2$ $3p^4$ → 12 + 6 → 16th group

N → [He] $2s^2 2p^3$ → 12 + 3 → 15th group

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③ Group no. of D-block elements

Grp no. → no. of $ns e^-$ + no. of ~~nde~~ $(n-1)d\ e^-$

eg. Cr - [Ar] $4s^1\ 3d^5$

group → 1+5 → 6th group.
→ D block.
→ 4th period.

eg 2. Fe → [Ar] $4s^2 3d^6$

group → 6+2 → 8th group.
→ D-block.
→ 4th period.

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

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

32 → ~~4s~~ [Ar] $4s^2 3d^{10} 4p^2$
→ P-Block
→ 4th period
→ 14th group

56 → [Xe] $6s^2$
→ S-Block
→ 6th period
→ 2nd group

2 → $1s^2$
→ S-Block
→ 1st period
→ ~~2nd~~ 18th group

27 → [Ar] $4s^2 3d^7$
→ D-Block
→ 4th period
→ 9th group

63 → [XE] $6s^2$ ~~5d~~ $4f^7$
→ F-block
→ 6th period
→ 39th group

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

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IUPAC Nomenclature of elements

(atomic num. > 100)

→ All names must end with "ium"

Bi + ium → Biium

Pent + ium → Pentium

Digit	name	Abrivation
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

Ex ① - Z=118 (UUO)
Ununoctium

② - Z= 104 (UUQ)
Unnilquadium

③ Z= 110 (UUn)
Ununnilium

④ Z=115 (UUP)
Ununpentium

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Diagnal 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 diagnal relationship.

2nd Period	Li	Be	B	C
3rd Period	Na	Mg	Al	Si

Li & Mg show similar properties.

eg. $Li_2CO_3 \xrightarrow{\Delta} Li_2O + CO_2$

$MgCO_3 \xrightarrow{\Delta} MgO + CO_2$

$Na_2CO_3 \xrightarrow{\Delta}$ No reaction

BeO, Al_2O_3 → Both are Amphoteric

Bridge Element

→ Elements Mg of 3rd period is bridge element

Be
↓
Mg → Ca, Sr, Ba
Mg → Zn, Cd, Hg

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

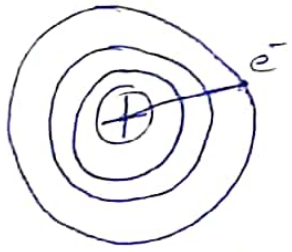
Periodic Properties

→ 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 - negitivity

→ Atomic Size / Atomic Radius

~~Atomic radius =~~



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

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

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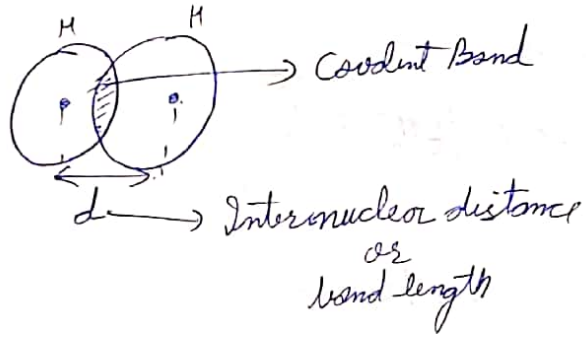
→ Since we are calculating the radius of atom in different bonded state (covalent bond → covalent radius → (for non-metal)
Metallic bond → Metallic radius → (for metal)
Vanderward's bond → Vanderward radius → (generally Nobel gases)
Ionic bond → Ionic radius

Covalent Radius →

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

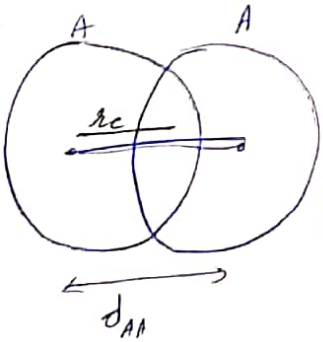
H_2



→ Bond length – It is ~~atoms~~ internuclear distance between two atoms in a bonded state.

→ Covalent radius.

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



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

$$R_{covalent} < R_{actual}$$

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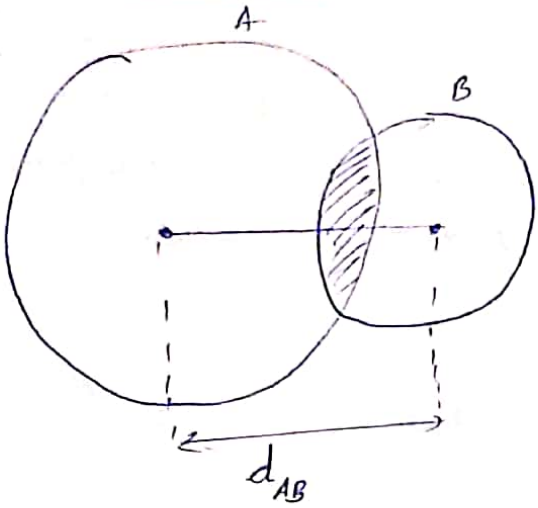
② Hetero-diatomic molecule (H-F, HCl)



Schoemaker - Stevenson Equation -

$$\cancel{d_{AB} = \gamma_A + \gamma_B - 0.09|EN_A - EN_B|}$$

~~γ = 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 Å

1 Å (Angstrom) = 10^{-10} m

1 pm (pico meter) = 10^{-12} m

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

1 μm (micro meter) = 10^{-6} m

1 mm (milli meter) = 10^{-3} m

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

(in pm)

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

$d_{HH} = 2A°$

$d_{HH} = R_H \times 2$

$2A° = R_H \times 2$

$\boxed{R_H = 1A°}$

$d_{ClCl} = 10A°$

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

$\boxed{R_{Cl} = 5\ A°}$

$d_{HCl} = 1 + 5 - 0.09\ (2.1 - 3)$

$= 6 - 0.09(0.9)$

$= 6 - 0.081$

$= 6.000$
$- 0.081$

$\boxed{5.919A°}$

Q2. Calculate bond length of AB if $R_A = 15$ Pm, $R_B = 20$ Pm, $EN_A = 2$, $EN_B = 1.5$

$d_{AB} = 15 + 20 - 9(2 - 1.5)$
$= 35 - 9(0.5)$

$= 35 - 4.5$

$\boxed{= 30.5\ pm}$

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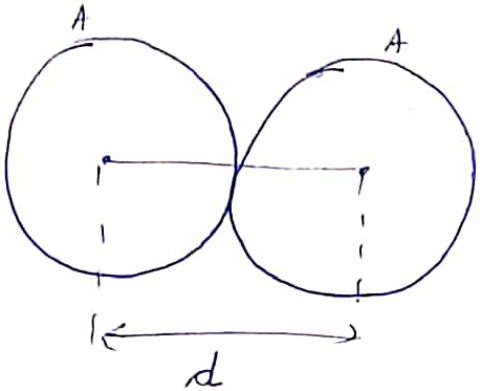
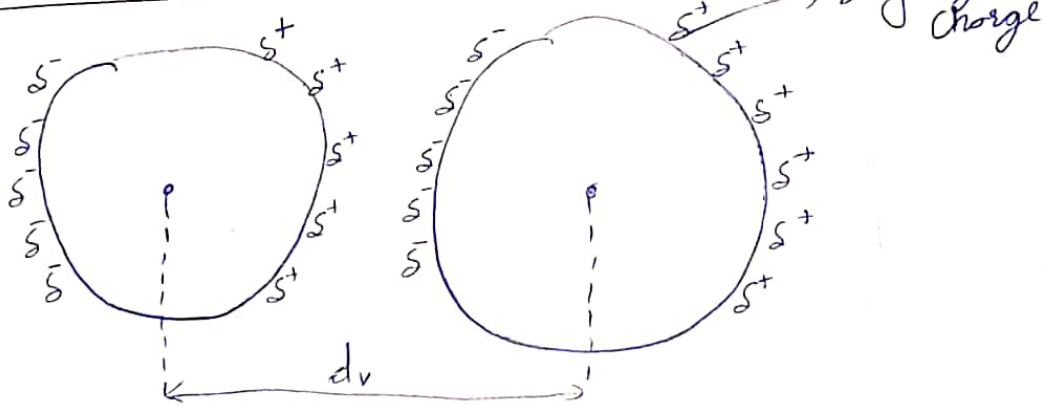
② (2) <u>Metallic Radius / Crystal Radius</u> – It is half the internuclear distance between two identical atoms bonded by a metallic bond.



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

$R_{metallic} \simeq R_{actual}$

③ <u>Vanderward's Radius</u> – due to weak forces



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

$R_v > R_{actual}$

$R_v > R_M > R_c$

R_v → vanderward, R_M → metallic, R_c → covalent

Scanned with CamScanner

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

factors affecting atomic size -

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

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

② $\boxed{\text{Shell no.} \propto \text{Size}}$

$n \uparrow$, Size $\uparrow$

Top across the period, Right

Left

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

$n \uparrow$

Size $\uparrow$ (in general)

bottom

for S or P block -

(In Period) -

1st period - H > He (expected)

He > H (Actual)

↓

Vanderwall radius

Scanned with CamScanner

2nd Period – Ne > Li > Be > B > C > N > O > F > Ne ← Noble gases

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

Ne → Vandervaal force

Li > Be → R_m

B > C > N > O > F → R_c

3rd Period – Ar > Na > Mg > Al > Si > P > S > Cl

Variation along the group –

G1	G 2	G 13	G14	G 15	G 16	G 17	G 18
							He
H						F	Ne
Li	Be	E	C	N	O	Cl	Ar
Na	Mg	x	Si	P	S	Br	Kr
K	Ca	C	Ge	As	Se	I	Xe
Rb	Sr	E	Sn	Sb	Te	At	Rn
Cs	Ba	P	Pb	Bi	Po		
Fr		I					
		O					
		N					

G-13 (Expected): B > Al > Ga > In > Tl

G-13 (Actual): B > Ga → due to d-orbital contraction; Ga = Al > In = Tl → due to f-block contraction or Lanthanides contraction

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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

Rade 1-4

Size of atom	$= 10^{-10}$ m	Approx
nucleus	$= 10^{-14}$ m	
p/n/e⁻	$= 10^{-15}$ m	

Note – Proton p^{+1}

β-particle → $^{0}_{-1}e$, e^{-1}, β^{-1}

α-particle → He^{2+}

Hydronium ion → H^{+} or H_3O^{+}

for-D-block –

Across the period:-

$\sigma \uparrow\uparrow$, $Z_{eff} \uparrow\uparrow$ (Both attraction force & repulsion force dominant.

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

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

Size ↓

(attraction force dominant)

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

(repulsion force dominant)

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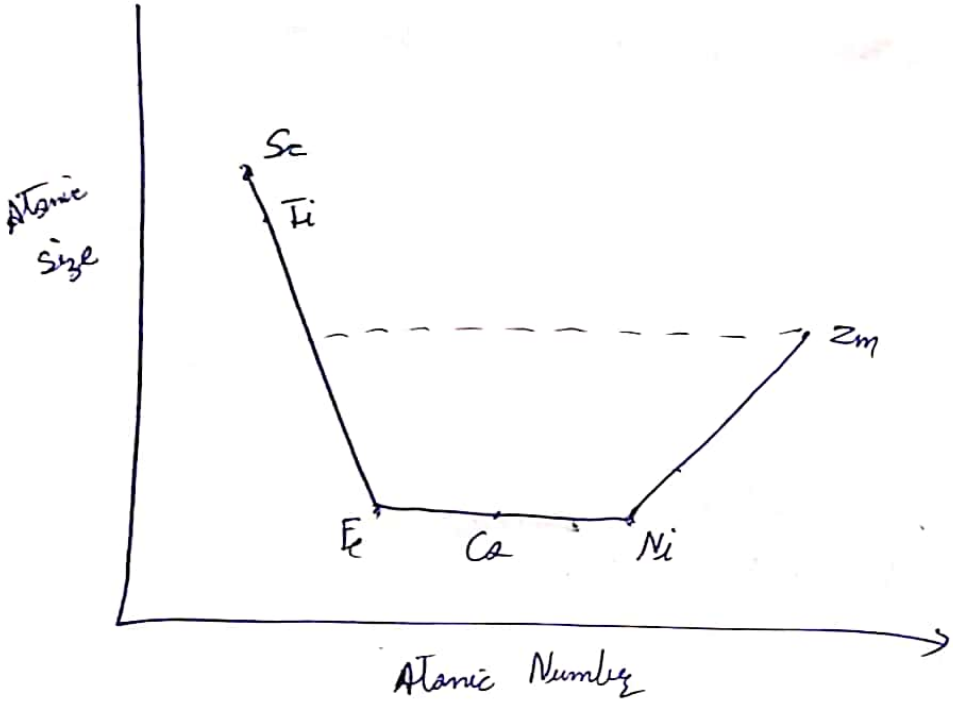

~~Group~~

variation along the groups

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

Sc < Y < La

Group-4 to 12 - (3d < 4d ≃ 5d)

G-3	G 4-12								
Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn
Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd
La	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg
	is	is	is	is	is	is	is	is	is → Almost equal

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Ionic Radius (Cations or Anions)

Cation -

$$A \longrightarrow A^{+} + e^{-}$$

Eg $P = 12$, $e = 12$ → $P = 12$, $e = 11$

Attraction F ↑, Size ↓

Size of atom > Size of Cation

Radius order :-

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

Eg $Na > Na^{+}$
$Mg > Mg^{2+}$
$Al > Al^{3+}$

Anion -

$$A + e^{-} \longrightarrow A^{-}$$

$P = 12$, $e = 12$ → $P = 12$, $e = 13$

Attraction ↓, Size ↑

Size of Atom < Size of Anion

Radius order

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

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Note :- Size

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

eg. $N < N^{3-}$

$O < O^{2-}$

$F < F^{-}$

<u>Size or Radius comparision between iso-electronic species.</u>

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

iso-electronic

$Z \uparrow$

σ = same

Z_{eff} ↑, Size ↓

eg2. $P^{3-} > S^{2-} > Cl^{-} > K^{+} > Ca^{2+}$

iso electronic

<u>Some Mixed Order.</u>

Li^{+} ← Mg^{2+} ← Be^{2+} → Al^{3+}

Na^{+} ← Ca^{2+}

K^{+} ← Sr^{2+}

Rb^{+} ← Ba^{2+}

Cs^{+}

increasing size →

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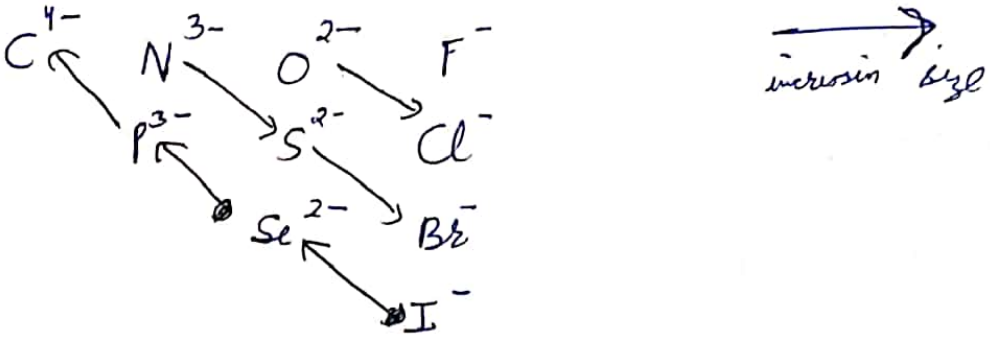

Q Radius Order?

① K, Rb, Cs

K < Rb < Cs ✓

② Be, F, O, N, Ne

Ne > Be > N > O > 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 ✓

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(8) Cu , Zn

$Cu < Zn$ ✓

(9) Fe , Fe^{2+} , Fe^{3+}

$Fe > Fe^{2+} > Fe^{3+}$ ✓

(10) O , O^{-1} , O^{2-}

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

(11) Pd, Ni , Pt

$Ni < Pd \simeq Pt$ ✓

(12) V, Ta, Nb

$V < Nb \simeq Ta$ ✓

(13) Sb, S, Se, Te

$S < Se < Te < Sb$ ✓

(14) Xe, Rn, Kr

$Kr < Xe < Rn$ ✓

(15) A^{+}, A^{2+}, A^{2-}, A^{-}, A

$A^{2-} > A^{-} > A > A^{+} > A^{2+}$ ✓

(16) H^{-}, Li^{+}

$H^{-} > Li^{+}$ ✓

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(17) F^-, Cl^-, Br^-, I^-

$F^- < Cl^- < Br^- < I^-$ ✓

(18) N^{3-}, S^{2-}, Br^-

$N^{3-} < S^{2-} < Br^-$ ✓

(19) Li^+, Mg^{2+}, Al^{3+}, Be^{2+}

$Be^{2+} < Al^{3+} < Mg^{2+} < Li^+$ ✓

(20) B, Al, Ga

$B < Ga \leq Al$ ✓

(21) Na, Cl

$Na > Cl$ ✓

(22) N^{3-}, Na^+

$N^{3-} > Na^+$ ✓

(23) Cl^-, Ca^{2+}

$Cl^- > Ca^{2+}$ ✓

(24) B, C

$B > C$ ✓

(25) Te, I

$Te > I$ ✓

(26) Mg, S

$Mg > S$ ✓

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(27) Cu, Cu^{+}, Cu^{2+}

$Cu > Cu^{+} > Cu^{2+}$ ✓

(28) Zn, Zn^{2+}

$Zn > Zn^{2+}$ ✓

(29) P, As

$P < As$ ✓

(30) Xe, Kr

$Kr < Xe$ ✓

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

$N^{3-} > N^{2-} > N^{+} > N^{2+}$ ✓

(32) Cr, Mo, W

$Cr < Mo \simeq W$ ✓

(33) Cu, Ag, Au

$Cu < Ag \simeq Au$ ✓

(34) Ni, Zn, Cu

$Ni < Cu < Zn$ ✓

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

$Cl^{-} > Cl > Cl^{+}$ ✓

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

High repulsion b/w of extra e^{-}

$\boxed{F^{-} < Cl^{-} < Br^{-} < I^{-} \quad H^{-} < Cl^{-}}$

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H.W. 29-04-2024

O-1 (Q1-Q14)

O-2 (Q1,2)

S-1 (Q13)

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

JA (2, 6)

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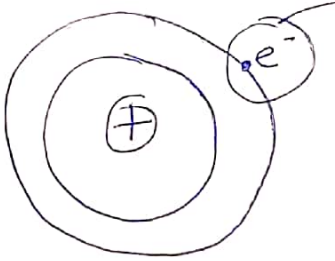
Note- Largest Cation $- Cs^+$
Smallest cation $- H^+$
Largest mono negative anion $= I^-$
Second - Largest mono negative anion $= H^-$
Smallest mono - negative anion $- F^-$

☆ Ionisation Energy / Ionisation Potential ☆
(I.E.) (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



energy absorbed [Tum mujhe energy do mai Tumhe e⁻ dunga]

$$A_{(g)} \xrightarrow{+I.E.} A^+_{(g)} + e^-$$

→ Unit of I.E. :- KJ/mole, e.v/atom, Kcal/mol, J/mol, Cal/mol
↓
electron volt

Enthalpy - It is the heat content of system.

Reactant E_R ⟶ Product E_P

ΔH (change in heat/enthalpy) $= E_P - E_R$

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$\Delta H = (+)ve \rightarrow$ ~~Endoth~~ Endothermic reaction $\rightarrow$ Energy absorbed

$\Delta H = (-)ve \rightarrow$ Exothermic Reaction $\rightarrow$ Energy ~~Released~~ Released.

$\Delta_{ai}H_{I.E}$ (Ionisation 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.

$$Na_{(g)} \longrightarrow Na^{+}_{(g)} + e^{-} \qquad \Delta H = 5.1 \text{ eV/atom} \quad (\Delta I.E)$$

$$Na^{+}_{(g)} + e^{-} \longrightarrow Na_{(g)} \qquad \Delta H = -5.1 \text{ eV/atom}$$

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

$$N_2 + 3H_2 \longrightarrow 2NH_3 \qquad \Delta H = x$$

$$\cancel{2N_2 + 6H_2 \longrightarrow} \qquad \Delta H' = 2x$$

$$2N_2 + 6H_2 \longrightarrow 4NH_3$$

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(3) If two or more than ~~Ab~~ two reactions are added, their ΔH are also added

$$Al_{(g)} \longrightarrow Al^{+}_{(g)} + e^- \qquad \Delta H_1 = (I.E_1)(Al)$$

$$Al^{+}_{(g)} \longrightarrow Al^{2+}_{(g)} + e^- \qquad \Delta H_2 = (I.E)_2 (Al)$$

$$Al^{2+}_{(g)} \longrightarrow Al^{3+}_{(g)} + e^- \qquad \Delta H_3 = (I.E)_3 (Al)$$

Adding the three

$$Al_{(g)} + \cancel{Al^{+}_{(g)}} + \cancel{Al^{2+}_{(g)}} \longrightarrow \cancel{Al^{+}_{(g)}} + \cancel{Al^{2+}_{(g)}} + Al^{3+}_{(g)} + 3e^-$$

$$Al_{(g)} \longrightarrow Al^{3+}_{(g)} + 3e^-$$

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

Successive Ionisation Energy:-

$$A_{(g)} \xrightarrow[-e^-]{(I.E)_1 (A)} A^{+}_{(g)} \xrightarrow[-e^-]{(I.E)_2 (A)} Al^{2+} \xrightarrow[-e^-]{(I.E)_3 (A)} Al^{3+}$$

$$(I.E)_1 (A) = (I.E)_1 (A^+)$$

$$(I.E)_3 (A) = (I.E)_2 (A^+) = (I.E)_1 (A^{2+})$$

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

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

$$B^{+6}_{(g)} \longrightarrow B^{+7}_{(g)} + e^-$$

A) $I.E_1$ (B)

2) $I.E_6$ (B)

3) $I.E_7$ (B) ✓

4) $I.E_5$ (B^{2+}) ✓

5) $I.E_4$ (B^{3+}) ✓

6) $I.E_3$ (B^{4+}) ✓

7) $I.E_1$ (B^{5+})

8) $I.E_1$ (B^{6+}) ✓

9) $I.E_2$ (B^{5+}) ✓

10) $I.E_6$ (B^{+1}) ✓

Q

Valence e^- calculation

~~$(I.E)_1$ (x).~~

Q. An element X, find valence shell e^-

$(I.E_1)_1$ (X) = 5 ev/atom

$(I.E_2)_2$ (x) = 13 ev/atom

$(I.E)_3$ (x) = 80 ev/atom

$(I.E)_4$ (x) = 100 ev/atom

sudden drastic change, means shell change

valence shell has $\boxed{2e^-}$

Q. Identify X

$(I.E)_1$ (x) = 8 ev/atom

$(I.E)_2$ (x) = 14 ev/atom

$(I.E)_3$ (x) = 25 ev/atom

$(I.E)_4$ (x) = 140 ev/atom

$(I.E)_5$ (x) = 172 ev/atom

$(I.E)_6$ (X) = 210 ev/atom

A) Be

B) Na

✓C) Al

D) B

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

$I.E_1 = 28$ ev/atom

$I.E_2 = 39$ ev/atom

$I.E._3 = 64$ ev/atom

$I.E._4 = 102$ ev/atom

$I.E._5 = 775$ ev/atom

$I.E._6 = 1020$ ev/atom

Valence Shell $e^- = 3$

factor's affecting I.E. →

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

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

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

② value of n :-

$n \uparrow$; $Size \uparrow$; $I.E. \downarrow$

$$I.E. \propto \frac{1}{Size} \propto \frac{1}{n}$$

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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 ionisation energy across period.

(generally)

Left $\xrightarrow{\text{Z eff} \uparrow;\ \text{Size} \downarrow;\ I.E \uparrow}$ Right

increase ionisation energy

Period 1 :- $H < He$ (max I.E is of He)

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

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

$2s^1$ $2s^2$ $2p^1$, $2p^2$ $2p^3$ $2p^4$ $2p^5$ $2p^6$ – fully filled

Penetration power

Half filled

$Li < B < Be < C < O < N < F < Ne$ (Actual)

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Reason :- ① Due to high penetration of 2s subshell compared to 2p subshell, the removal of e^- from 2s of ~~Be~~ 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 requirment 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}.

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Period-3 → Na < Mg < Al < Si < P < S < Cl < Ar

$3s^1$ (Expected) $3s^2$ $3p^1$ $3p^2$ $3p^3$ $3p^4$ $3p^5$ $3p^6$

Na < Al < Mg < Si < S < P < Cl < Ar (Actual)

Q1. (I.E.) order

Li^+ Be^+ B^+ C^+ N^+ O^+ F^+ Ne^+

$1s^2$ $2s^1$ $2s^2$ $2p^1$ $2p^2$ $2p^3$ $2p^4$ $2p^5$ (removed $1e^-$)

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

Note → ① In periodic table He has highest Ionisation Energy.

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

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

Q2. I.E$_3$ Mn I.E.$_3$ Fe

Mn^{2+} $3d^3$ $4s^0\ 3d^5$ → Half filled

> I.E$_3$

Fe^{3+} $3d^4$ $4s^0\ 3d^6$

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Q. I.E.
Q3.

Group 11		Group 12
Cu $3d^{10} ns^1$	<	Zn $3d^{10} ns^2$
Ag $4d^{10} ns^1$	<	Cd $4d^{10} ns^2$
Au $5d^{10} ns^1$	<	Hg $4d^{10} ns^2$

Variation of I.E. down the group :-

Top

$n \uparrow$
Size $\uparrow$
I.E. $\downarrow$

[in general]

Bottom

G1	G 2	G 13	G 14	G 15	G 16	G 17	G 18
Li	Be	E	C	N	O	F	Ne
✓	✓		✓	✓	✓	✓	✓
Na	Mg	X	Si	P	S	Cl	Ar
✓	✓		✓	✓	✓	✓	✓
K	Ca	E	Ge	As	Se	Br	Kr
✓			✓	✓	✓	✓	✓
Rb	✓	C	Sn Pb	Sb	Te	I	Xe
✓	Sr	T		✓	✓	✓	✓
Cs	✓	I	✓			At	Rn
	Ba	R	Sn	Bi	Po		

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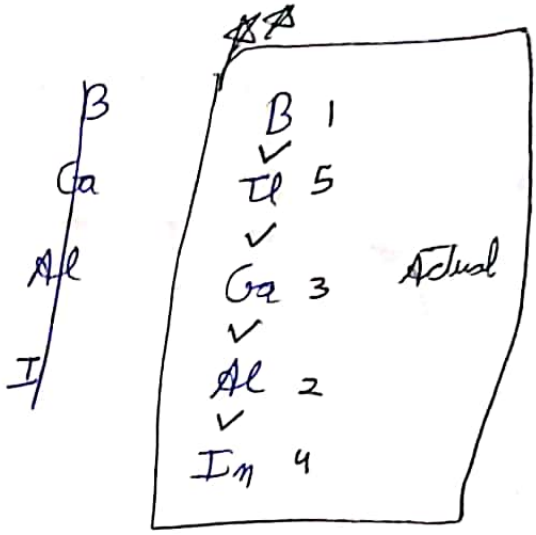
Exception 1 - G-13

Expected

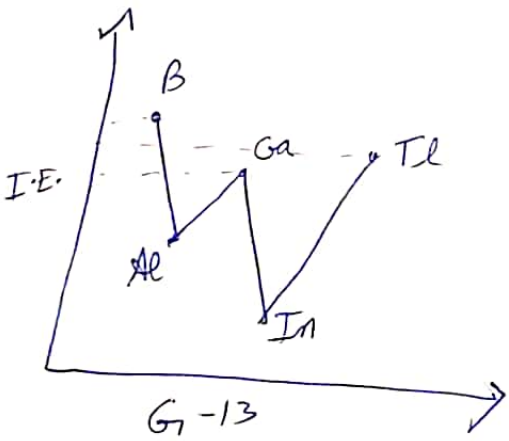


Actual

Exception 2 - G-14

C > Si > Ge > ~~Pb~~ Sn > Pb (Expected)

C > Si > Ge > ~~Sn~~ > Pb > Sn (actual)



Explanation (~~Exception 1~~) - In Boron family, Due to lanthanides and D-~~orbital~~ contraction in Tl, its Z-eff increases sharply due to which Ionisation Energy also increases and becames next to boron.

Similar effect is observed with respect to Tin & Lead (Sn, Pb) but Ionisation energy of lead does not increases too much like thalium.

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D-Block

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

Iso - Protonic :-

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

Z same e^- ↓

~~Zeff same~~ σ ↓

~~σ~~ Zeff ↑

I.E. ↑

eg 1. $Fe < Fe^{2+} < Fe^{3+}$

Iso Electronic :-

$${}_{3}Li^{+} > {}_{2}He$$

$e^- \rightarrow$ same

$\sigma \rightarrow$ same

Z ↓ Zeff ↓ I.E ↓

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Q① I.E order

① H^- Li^+ Be^{2+}

$H^- < Li^+ < Be^{2+}$

② Sn^{2+} Sn^{4+}

$Sn^{2+} < Sn^{4+}$

③ O^{2-} O^- O

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

④ C^{2+} C^{4+}

$C^{2+} < C^{4+}$

⑤ Ne Na^+

$Ne < Na^+$

⑥ Na Na^+

$Na < Na^+$

⑦ F F^-

$F^- < F$

⑧ Na^+ Mg^{2+} Al^{3+}

$Na^+ < Mg^{2+} < Al^{3+}$

⑨ O^{2-} F^- Na^+ Mg^{2+}

$O^{2-} < F^- < Na^+ < Mg^{2+}$

~~Li Be~~

Q2:

① Li Be

$Li < Be$

② Be B

$B < Be$

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(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

Scanned with CamScanner

(14) Ge Mn

Ge < Mn ✓

(15) V Nb Ta

Ta > Nb ≃ V ✓

(16) Sc Y La

Sc > Y > La ✓

(17) H^- Li^+ Be^{2+}

$H^- < Li^+ < Be^{2+}$ ✓

Application of I.E —

(1) As ionisation energy increases, metallic character decreases

I.E.↑, M.C.↓

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

Metallic character order

Li < Na < K < Rb < Cs

I.E.↓, M.C.↑

Eg 2. Be < B
2S 2P

Eg 3. Be < Mg

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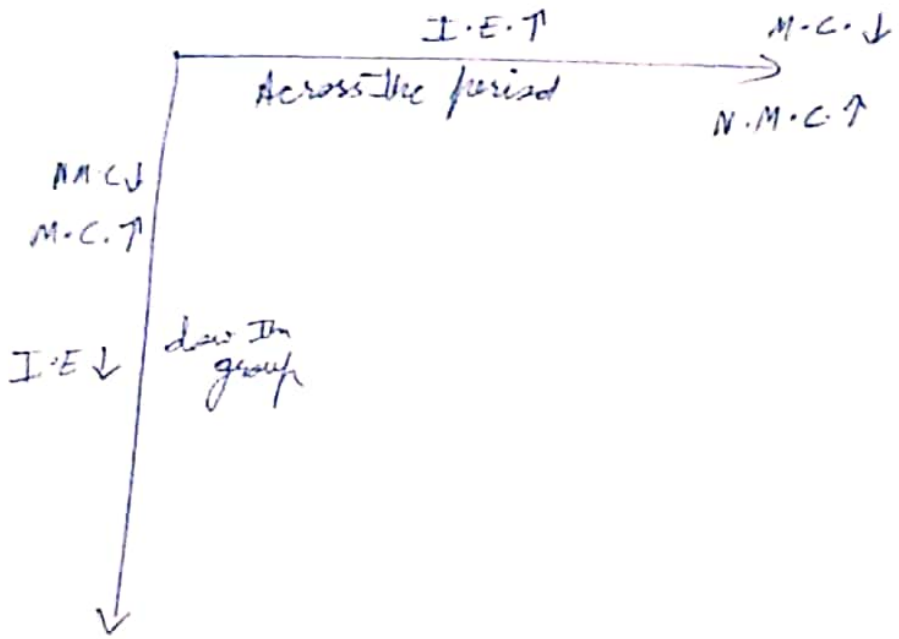
(2) As the ionisation energy increases, non-metallic character increases.

I.E ↑ N.M.C ↑

Eg 1. F > Cl > Br > I

Eg 2 O > S > Se > Te



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

Eg. Li < Na < K < Rb < Cs

Be < Mg < Ca < Sr

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(4) Oxidation → It means removal of e^-.

$$Na_{(g)} \longrightarrow Na^+_{(g)} + e^-$$

Reduction → Addition of e^-.

$$Cl_{2(g)} + e^- \rightarrow \frac{1}{2} Cl^-_{2(g)}$$

~~Oxidising Power~~ →

<u>Oxidising Power</u> → It means the tendency of an element to oxidise other elements by reducing itself.

<u>Reducing Power</u> → Tendency of an ~~element~~ element to reduce other elements by oxidising itself. (Oxidise easily hogo)

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

<u>I.E ↓ Reducing Power ↑</u>

(5) a) If

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

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

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Na		eV/atom
Na^{+}	IE_1	05.1
Na^{2+}	IE_2	41.2
	$I.E_3$	71.62.

36.1 Na^{+} more stable

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

then higher oxidation will be more stable

Eg2. $(I.E)_2 - (I.E)_1 = 7.4$ eV/atom < 11 eV/atom

(Mg)

Mg^{2+} is more stable than Mg^{+}

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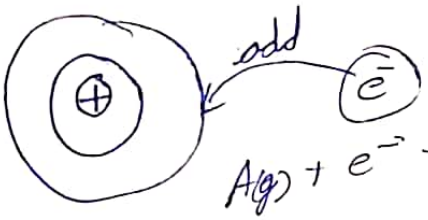
Electron Affinity (E.A.) – Amount of energy released when at electron is added to the valence shell of an isolated gasious atom or ion.



$A_{(g)} + e^- \rightarrow A^-_{(g)}$

let energy released = 100 kJ/mol

E.A = 100 kJ/mol

$\Delta H_{E.g}$ (electron gain enthalpy) = −100 kJ/mol (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 gasious atoms.

$B_{(g)} + e^- \rightarrow B^-_{(g)}$

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

energy supplied = 100 kJ/mol

$\Delta H_{Eg} = +100$ kJ/mol

⇓

endothermic

(e^- zabardasti diya)

at 0K (0 kelvin)
$\Delta H_{E.G.} = -EA$

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Note:- ① $\Delta H_{Eg} = \ominus ve$; $E \cdot A = \oplus ve \Rightarrow$ energy released

② $\Delta H_{Eg} = \oplus ve$; $E \cdot A = \ominus ve$ or zero $\Rightarrow$ energy absorbed

Note:-

① - In case of Inert ~~Gas~~ Gas or some stable configuration energy is ~~does~~ not released, instead of ~~these~~ this, we have to supply energy.

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

$\Delta H_{eg_1} = \oplus ve$

$E \cdot A_1 = \ominus ve$

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

$\Delta H_{eg_1} - \ominus ve$

$E \cdot A_1 - \oplus ve$

* $A_{(g)} + e^- \rightarrow A^-_{(g)}$ ΔH_{Eg_1} or EA_1

$A^-_{(g)} + e^- \rightarrow A^{2-}_{(g)}$ ΔH_{Eg_2} or EA_2

$A^{2-}_{(g)} + e^- \rightarrow A^{3-}_{(g)}$ ΔH_{Eg_3} or EA_3

For Any Atom $\rightarrow$ $\boxed{\Delta H_{EG_2} = \oplus ve \text{ (always)}}$

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

$\Delta H \Rightarrow$ supplied

$$\left[\Delta H_{E.G.E_3} > \Delta H_{EGE_2} > \Delta H_{EGE_1}\right]$$

ΔH_{EGE_3} → ⊕ve

ΔH_{EGE_1} → ⊕ve or ⊖ve

Note:- EA_1 can be positive or negative

↓

(Be, Mg, N, All inert gases)

ΔH_{EGE_1} can be ⊖ve or ⊕ve

$A_{(g)} + e^- \rightarrow A^-_{(g)}$ $\quad \Delta H_{EGE_1} \Rightarrow$ ⊕ve / ⊖ve

$A^-_{(g)} + e^- \rightarrow A^{2-}_{(g)}$ $\quad \Delta H_{EGE_2} \Rightarrow$ ⊕ve always

$A_{(g)} + 2e^- \rightarrow A^{2-}_{(g)}$ $\quad \Delta H_{EGE_{Total}} = \Delta H_{EGE_1} + \Delta H_{EGE_2}$

⊕ve/⊖ve + ⊕ve

⊕ve always

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For Cations :-

$$A^+_{(g)} + e^- \longrightarrow A_{(g)}$$

energy released

$|EA_1|\ (A^+)$

or

$|\Delta H_{EGE_1}|\ (A^+)$

$$A^+_{(g)} \longrightarrow A^+_{(g)} - e^-$$

energy gain

$|I.E_1|\ (A_{(g)})$

$$|\Delta H_{EGE_1}|(A^+) = |\Delta H_{IE_1}|\ (A)$$

$$|EA_1|\ A^+ = |I.E_1|\ A$$

For Anion :-

$$A_{(g)} + e^- \longrightarrow A^-_{(g)}$$

$|\Delta H_{EGE_1}|\ (A)$

$|I.E_1|\ (A^-)$

$$A^-_{(g)} \longrightarrow A_{(g)} - e^-$$

$$|\Delta H_{EGE_1}|_{(A_{(g)})} = |I.E_1|\ (A^-_{(g)})$$

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~~Trends in E~~

Trends in E.G.

① Zeff :- Zeff ↑ Attraction ↑ E.A ↑

② value of n :- n ↑ ; Size ↑ ; Zeff ↓ ; E.A ↓

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

EA: ns > np > nd > nf

जो nucleus के पास होगा electron affinity ज्यादा होगी।

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

In Periodic Table

Zeff ↑; E.A ↑ (in general) →

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

E.A. ⊖ve

Ne < Be < N < B < Li < C < O < F (Expected)

Zeff

Penetration power

Period-1 H > He

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

$\ominus$ve Zeff

~~$Ar < Mg < P < Al < Na < Si < S < Cl$ [Actual]~~

Zeff Penetration Power

Si ek e- lekhne fill ho jayega Half

$\ominus$ve Si ek e- lekhne fill ho jayega Half

$Ar < Mg < Al < Na < P < Si < S < Cl$ [Actual]

Penetration power Zeff

exception

Top

$n\uparrow$; $EA\downarrow$

Bottom

S-Block: $Li > Na > K > Rb > Cs$

$Be > Mg > Ca > Sr$

2-nd Period :-	B	C	N	O	F
	^	^	^	^	^
3-rd Period :-	Al	Si	P	S	Cl

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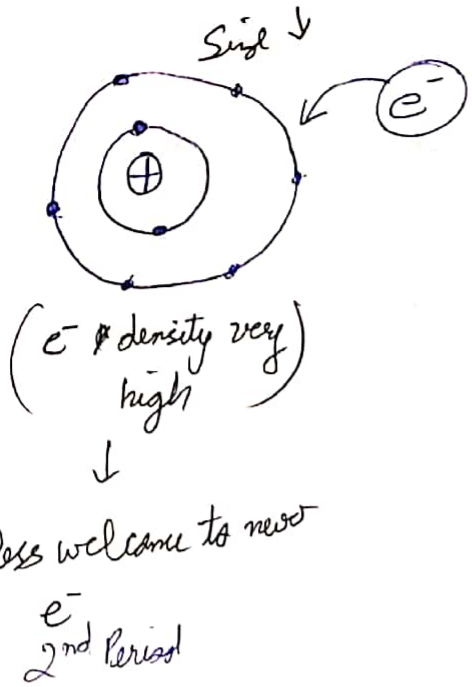
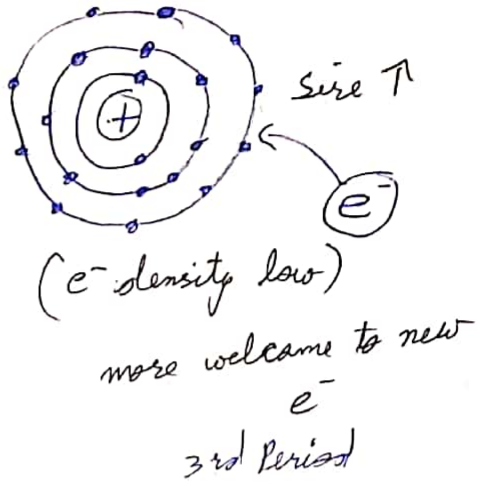
Reason:-





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

Halogen Family -

$Cl > F > Br > I$

Oxygen Family:-

$S > Se > Te > O$

Nitrogen:-

$N < P < As < Sb$

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

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Note:-

① Oxygen and Nitrogen shows least e^- affinity in their respective groups.

② ~~Cl~~ Chlorine has highest e^- affinity in periodic table and florine comes to 2nd position.

③ As E.A. ↑ ; Non metallic character ↑

Q1. -E A. order?

① Cl , Br , I

$I < Br < Cl$ ✓

② Na, Al, Mg, Si

$Mg < Al < Na < Si$ ✓

③ F , Cl

$F < Cl$ ✓

④ O , S

$O < S$ ✓

⑤ C, N

$N < C$ ✓

⑥ Se, S

$Se < S$ ✓

⑦ P, As, Sb

$P < As < Sb$ ✓

⑧ Li , Be , B

$Be < B < Li$ ✓

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(9) O, F,

O < F ✓

(10) O, F, Cl

O < F < Cl ✓

(11) N, O, F

N < O < F ✓

(12) C, O

C < O ✓

(13) O, N

N < O ✓

Q 2. write the sign of ΔH in the following.

(1) $P + e^- \longrightarrow P^{\ominus}$ $\Delta H = \ominus ve$ ✓

(2) $Ca + e^- \longrightarrow Ca^-$ $\Delta H = \ominus ve$ ✓

(3) $N + e^- \longrightarrow N^-$ $\Delta H = \oplus ve$ ✓

(4) $He + e^- \longrightarrow He^-$ $\Delta H = \oplus ve$ ✓

(5) $Na^+ + e^- \longrightarrow Na$ $\Delta H = \ominus ve$ ✓

(6) $Mg^{2+} + e^- \longrightarrow Mg^+$ $\Delta H = \ominus ve$ ✓

(7) $Kr + e^- \rightarrow Kr^-$ $\Delta H = \oplus ve$ ✓

(8) $N^- + e^- \rightarrow N^{2-}$ $\Delta H = \oplus ve$ ✓

(9) $S \rightarrow S^{\oplus} + e^-$ $\Delta H = \oplus ve$ ✓

(10) $Cl \rightarrow Cl^{\oplus} + e^-$ $\Delta H = \oplus ve$ ✓

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Application of E.A.

E.A means ↑ → Gain of e^- easily → Reduce easily
↓
Oxidise rapidly
↓
Oxidising Power ↑

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

Q Compare I.E. of $F^{\ominus}$ and $Cl^{\ominus}$

$F + e^- \rightarrow F^{a-}$ $\quad |E.A|_F$

$Cl + e^- \rightarrow Cl^-$ $\quad |E.A|_{Cl}$

$F^- \longrightarrow F + e^-$ $\quad |I.A|_F$

$Cl^- \longrightarrow Cl + e^-$ $\quad |I.A|_{Cl}$

$E.A = I.A$

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

so

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

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Q2. Compare I.E. order in following.

① C^- O^-

$C^- < O^-$ ✓

② N O

$N > O$ ✓

③ Ne^- Be^-

$Ne^- < Be^-$ ✓

④ O^- S^-

$O^- < S^-$ ✓

⑤ Ne Be

$Ne > Be$ ✓

⑥ Al^- B^-

$Al^- > B^-$ ✓

Q3. Compare I.E.

F^- Cl^- F Cl

~~$Cl > F > Cl^- > F^-$~~

~~$F > Cl > Cl^- > F^-$~~

~~$Cl^- > F^- > F > Cl$~~

$F > Cl > Cl^- > F^-$

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Q 34. I.E.

$O \quad S \quad O^- \quad S^-$

$O > S > S^- > O^-$

Q S.E.A. order. ?

$F \quad F^+ \quad Cl \quad Cl^+$

~~$Cl^+ > F^+ > Cl > F$~~

~~$F^+ > Cl^+ > Cl > F$~~

~~$Cl^+ > F^+ > Cl > F$~~

$\boxed{F^+ > Cl^+ > Cl > F}$

Q $O \quad O^+ \quad S \quad S^+$

$O^+ > S^+ > S > O$

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O-1 (Q 15-27)

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

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

~~(S-2) (Q 8, 5)~~

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

J-M (Q 1, 4, 5, 8, 10, 11, 14, 19, 21, 26, 26, 29, 33, 36, 38)

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Electron Negativity (E.N) (χ)

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

$\overset{\delta^+}{H}$ ⊙ → F δ^- (partial ⊖ve charge)

↳ shared Pair of e^- moves slightly to F as it is powerful.

dipole (as two poles ⊕ve & ⊖ve

H — ⊙ — H {No Dipole}

→ Electron Negativity has no unit.

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

① Pauling Scale of E.N. :-

→ It is a relative scale of Electron Negativity (Reference element – Flourine (F)).

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

$$\boxed{(EN_B - EN_A) = 0.208\sqrt{E_{A-B} - \sqrt{E_{A-A} \times E_{B-B}}}}$$

ΔEN

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$EN_A \Rightarrow$ Electron Negativity of A

$EN_B \Rightarrow$ Electron Negativity of B

$E_{A-B} \Rightarrow$ Bond energy of A-B bond (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)

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

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

Q1. If $EN_A = 2$

$EN_B > EN_A$

$E_{A-B} = 76$ KCal/mol

$E_{A-A} = 81$ KCal/mol

$E_{B-B} = 64$ KCal/mol

$EN_B = ?$

$EN_B = 2 + 0.208\sqrt{76 - \sqrt{81 \times 64}}$

$= 2 + 0.208\sqrt{76 - 72}$

$= 2 + 0.208 \times 2$

$= 2 + 0.416$

$$\boxed{EN_B = 2.416}$$

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Q2. Calculate EN_A if

$EN_B = 3$

$E_{AB} = 65$ Kcal/mol

$E_{A-A} = 49$ Kcal/mol

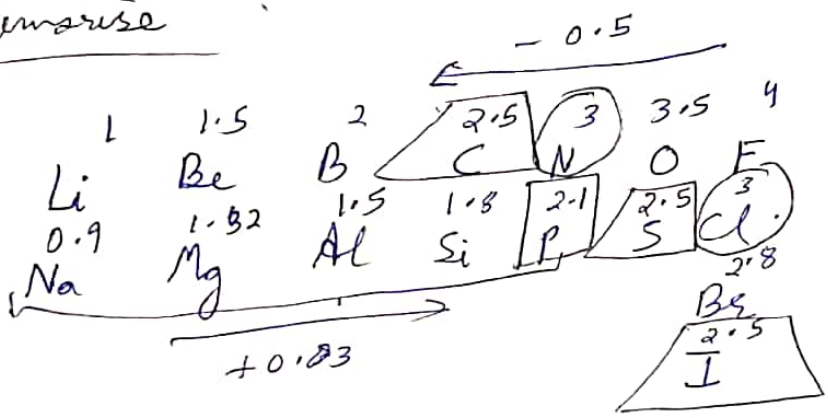
$E_{B-B} = 64$ Kcal/mol

$EN_A = 3 + 0.208 \times 3$

$= 3 + 0.624$

$\boxed{= 3.624}$

Memorise



$\boxed{M \rightarrow 2.1}$

Note- $C \simeq I \simeq S = 2.5$

$M \simeq P = 2.1$

$N \simeq Cl = 3$

$F = 4$ (Max)

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② Mulliken Scale

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

I.E, E.N in eV/atom.

$$\boxed{(E.N)_P = \frac{(E.N)_M}{2.8}}$$

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

$I.E(Cl^-) = 4\ eV/atom = |E.A|\ Cl$

$E.A(Cl^+) = 13\ eV/atom = |I.E|\ Cl$

$$E.N.(M) = \frac{13+4}{2}$$

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

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

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

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③ Allred Rochow Scale :-

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

Z_{eff} → effective nuclear charge

r → covalent radius in Å°

$1\ Å = 10^{-10}\ m$

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

$$E.N._p = E.N._{AR} + 0.744$$

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

Factors Affecting Electron Negativity

① **Z_{eff} :-**

$Z_{eff}\uparrow$ $\quad$ $E.N\uparrow$

② **Value of n :-**

$n\uparrow$ $\quad$ $E.N\downarrow$

③ **Charge :- (+ve)**

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

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

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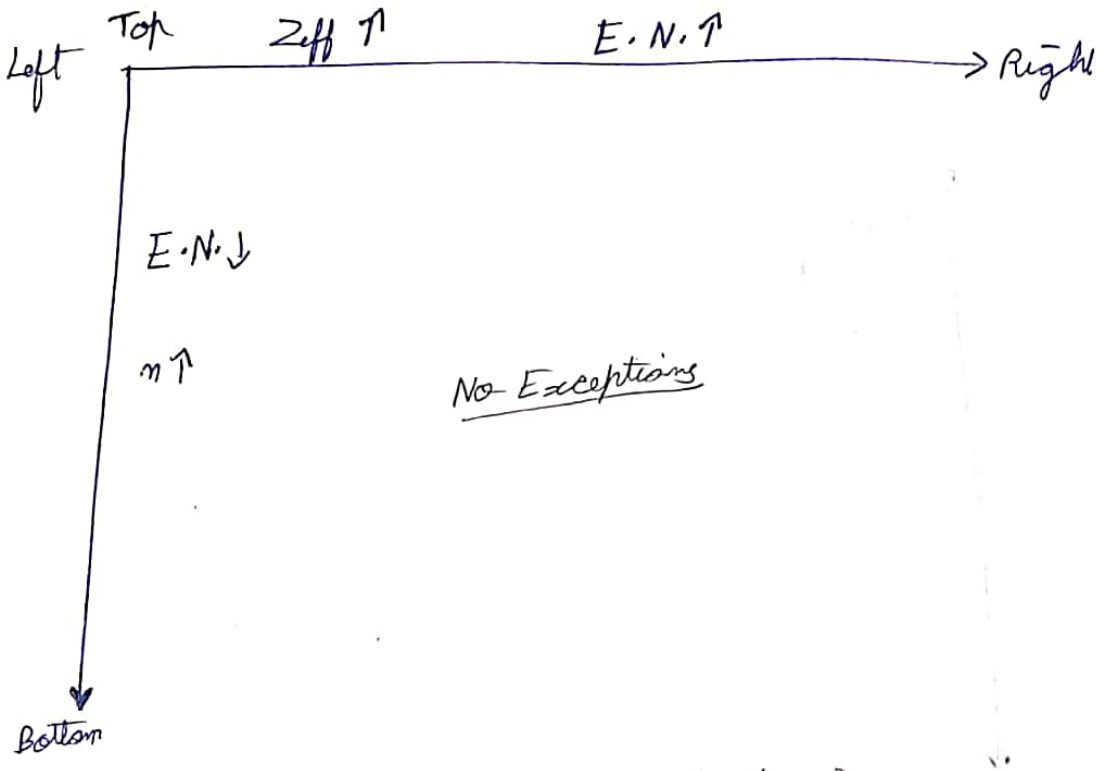
~~Note~~ -

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

Variation across Periodic Table



No Exceptions

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

① $C - H$
$H^{\delta+}$ $C^{\delta-}$

② $H - F$
$H^{\delta+}$ $F^{\delta-}$

③ $O - H$
$O^{\delta-}$ $H^{\delta+}$

④ $N - F$
$N^{\delta+}$ $F^{\delta-}$

Polar

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(a) F - F

(s) Non - Polar

Applications of Electron Negativity

→ (1)

$$\xrightarrow[M.C.\downarrow;\ N.M.C.\uparrow]{L\ \text{size}\downarrow;\ Z_{eff}\uparrow;\ EN\uparrow;\ I.E\uparrow;\ EA\uparrow} R$$

Top

$E.N\downarrow$

$N.M.C\downarrow$

$M.C\uparrow$

↓

Bottom

(2) If $\Delta EN \neq 0$ {Polar} eg H-F

$\Delta EN = 0$ {Non-Polar} eg. F-F

(3) Prediction of Bond Length of Heterodiatomic molecule.

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

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

$$\boxed{\% \text{ Ionic Character} = 16|\Delta EN| + 3.5(\Delta EN)^2}$$

ΔEN in Pauling Scale

$EN \uparrow$; % I.C $\uparrow$; % covalent Character $\downarrow$

% I.C + % 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.

$EN_{Cl} = 3$

$EN_{F} = 4$

$\Delta EN = 4 - 3 = 1$

$IOC = 16 + 3.5$

$\boxed{IC = 19.5}$

$CC = 100 - IC$

$\boxed{CC = 80.5}$

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Q3. I.C% in HCl

$\Delta EN = 2.1 - 3$

$= 0.9$

$I.C = 16(0.9) + (0.9)^2$

$= 14.4 + 0.81 \times 3.5$

$\cancel{= 22.5}$ $= 14.4 + 2.8$

$= 17.2$

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

Q4. ΔEN for $I.C = 50$

$\Delta EN = x$

$50 = 16x + 3.5(x^2)$

$7x^2 + 32x - 100$

$\cancel{x = -14 \pm \sqrt{196 + 2800}}$

$\cancel{= -14 \pm \sqrt{2996}}$

$\cancel{= -14 \pm 54.}$

$x = -32 \pm \sqrt{1024 + 2800}$

$= -32 \pm \sqrt{3824}$

$= \frac{-32 \pm 61.}{14}$

$= \frac{-16 \pm 30.5}{7}$

$= \frac{14.5}{7}$

$= 2.1$

$$\begin{array}{r} 496 \\ 416 \\ \hline 104 \\ 224 \\ 122 \end{array}$$

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

$\Delta EN = 2.1$

$\% I.C = 50\%$

$\% C.C = 50\%$

Predominantly ionic compound

Nature of compound → ionic

$\Delta EN < 2.1$

$\% I.C < 50\%$ $\quad \% C.C > 50\%$.

covalent compound

$\Delta EN > 2.1$

$\% I.C > 50\%$. $\quad \% C.C < 50\%$.

ionic compound

Q1. If the value of $\sqrt{E_{A-B} - \sqrt{A-A \times E_{BB}}}$ is termed as Δ, if Δ is given as 14.4 for elements A & B and bond length of A A bond is 1 A° and bond length of B-B bond is 1.2 A°

Calculate –

a) Calculate I.C of A-B bond.

b) Covalent bond length of A-B bond.

energies in KJ/mol

$\Delta EN = 0.102 \times 14.4$

$= 1.4688$

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

$\% I.C = 23.5008 + 0.7556$

$\boxed{= 24.25\%}$

$$\begin{array}{r} 144 \\ 102 \\ \hline 288 \\ 14400 \\ \hline 14688 \end{array}$$

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Bond length = $D_{AB} = 0.5 + 0.6 - 0.09(1.4688)$

$= 0.5 + 0.6 - 0.132$

$\boxed{= 0.968 \text{ A}^\circ}$

1.40

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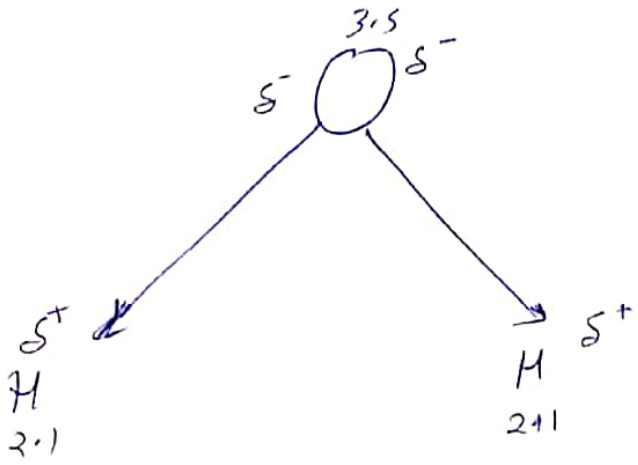
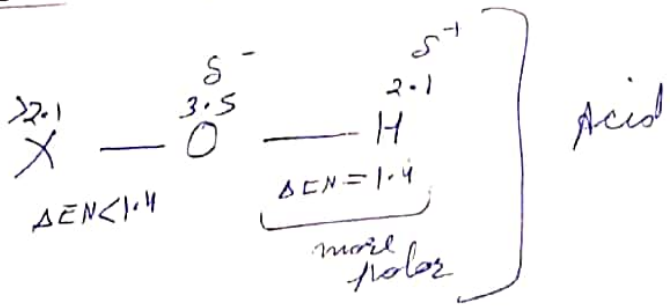
Nature of Hydroxide

→



more polar bond will break in aqueous solution.

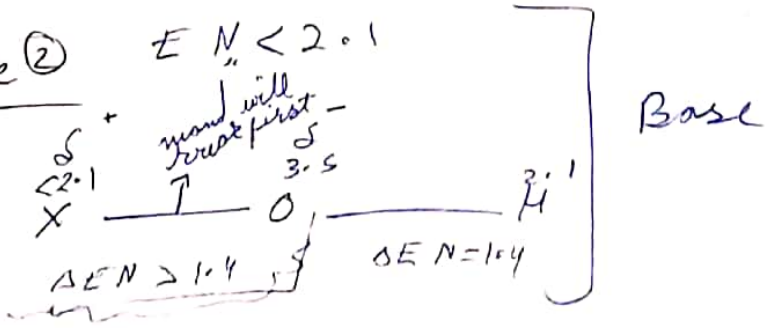
Case ① $EN_x > 2.1$



$X-O^{\ominus}$ + $H^{\oplus}$

Conjugate Base

Case ② $EN < 2.1$



$X^{\oplus}$ + $OH^{\ominus}$

↓

Conjugate acid

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Q find Acid & Base.

① $\overset{1.5}{A}$ — O — H (Base)

② $\overset{2.4}{B}$ O H
Acid

③ $\overset{1.7}{C}$ OH
Base

④ $\overset{2.7}{D}$ OH
Acid

Acidic And Basic Strength of Hydroacides.

$$\underset{\text{Acid}}{AOH} \longrightarrow \underset{\text{conjugate Base}}{AO^-} + H^+$$

$$\underset{\text{Base}}{XOH} \longrightarrow \underset{\text{conjugate acid}}{AX^+} + OH^-$$

$$\xleftarrow[\text{Basic Strength} \uparrow]{\text{E.N.} \downarrow} \boxed{2.1} \xrightarrow[\text{Acidic Strength} \uparrow]{\text{E.N.} \uparrow}$$

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Ex:. Acidic strength

BOH < DOH

$\overset{2.4}{B}$ —— $\overset{3.5}{O}$ —— $\overset{2.1}{H}$

$\Delta EN = 1.1$ $\Delta EN = 1.4$

$\overset{2.7}{D}$ —— $\overset{3.5}{O}$ —— $\overset{2.1}{H}$

$\Delta EN = 0.8$ $\Delta EN = 1.4$

↓

less ΔEN, less strength in bond

break easily

↓

Conjugate Base more stable

↓

Thus, Acid strength ↑

Basic strength

Ex(2) $\overset{1.5}{AOH}$ > $\overset{1.7}{COH}$

A —— $\overset{3.5}{O}$ —— H

$\Delta EN = 2$ $\Delta EN = 1.4$

C —— $\overset{3.5}{O}$ —— H

$\Delta EN = 1.8$ $\Delta EN = 1.4$

↓

Break easily (because of high polarity)

↓

More Basic strength ↑

↓

More stable Conjugate Acid.

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Bond Strength / Bond Energy

Bond strength $\propto$ ~~EN~~ ΔEN

Ex- H — F (2.1, 4) H — Cl (2.1, 3)

H — I (2.1, 2.5) H — Br (2.1, 2.8)

HF > HCl > HBr > HI (Bond strength)

$\Delta EN \uparrow$, bond strong $\uparrow$

HF < HCl < HBr < HI (Acidic strength)

$\Delta EN \uparrow$, Acidic strength $\downarrow$

Types of Oxides

① Basic oxide - (M B) → Basic (M ↑ metal)

→ usually metals form basic oxides which reacts with acid.

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

Ex:- $Na_2O + H_2O \longrightarrow 2NaOH$

$MgO + H_2O \longrightarrow Mg(OH)_2$

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M.N. 10-05-2024

O-1, S-1 (complete)

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② - Acidic Oxides (NMA) → Acidic

(NM → non metals)

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

→ when it dissolves in water forms acids

eg. $CO_2 + H_2O \rightarrow H_2CO_3$

$SO_3 + H_2O \rightarrow H_2SO_4$

$N_2O_5 + H_2O \rightarrow 2HNO_3$

③ - Amphoteric Oxides -

→ They react with both acids and bases.

Ex:-

ZnO	Al_2O_3	BeO	Ga_2O_3
जनाब	अली ने	बेकार	गाया
	SnO		
PbO	SnO_2	Cr_2O_3	
PbO_2		Cried.	
पंजाबी	Sang		

MnO_2, As_2O_3, Sb_2O_3, V_2O_5

④ - Neutral Oxides

→ ~~N_2O, NO~~

→ Oxides which do not react with Acid as well as base.

→ Ex. N_2O, NO, CO, H_2O → (If treated as solvent, then considered neutral)

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

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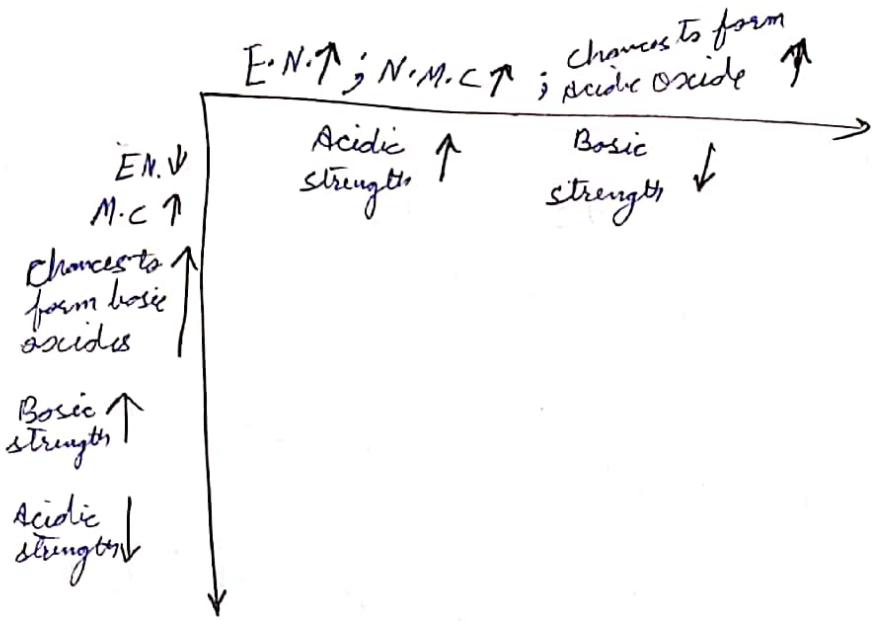
Periodicity of oxides



Q Acidic strength order.

① $Li_2O < BeO < Be_2O_3 < CO_2 < N_2O_5$ EN↑ ; A.S↑

② $\underbrace{Na_2O \quad MgO}_{\text{Basic oxides}}$ $\underbrace{Al_2O_3}_{\text{Amphoteric}}$ $\underbrace{SiO_2 \quad P_2O_5 \quad SO_3 \quad Cl_2O_7}_{\text{Acidic oxides}}$

$Na_2O < MgO < Al_2O_3 < SiO_2 < P_2O_5 < SO_3 < Cl_2O_7$

③ $Li_2O \quad Na_2O \quad K_2O \quad Rb_2O \quad Cs_2O$ EN↓ ; Acidic strength ↓

$Li_2O > Na_2O > K_2O > Rb_2O > Cs_2O$

④ $CO_2 \quad SiO_2 \quad SnO_2 \quad PbO_2$ (Basic strength order)

$CO_2 < SiO_2 < SnO_2 < PbO_2$

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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, Na, K, Rb, Cs, Fr)

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

In Bonded state

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

→ Hydride with non metal/other elements ⇒ +1 {NH_3, PH_3, CH_4}

③ Oxygen general oxidation state ⇒ −2

H_2O_2 ⇒ −1
KO_2 ⇒ −½
KO_3 ⇒ −⅓
OF_2 ⇒ +2
O_2F_2 ⇒ +1

④ Flourine ⇒ −1 (Always)

Acidic Strength ∝ oxidation state

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Q. ① CO_2 (find for carbon)

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

-2 → general oxidation state of oxygen

0 → charge on whole compound

2 → no. of oxygen

$x - 4 = 0$

$\boxed{x = 4}$

② SO_2

$x + 2(-2) = 0$

$x - 4 = 0$

$\boxed{x = 4}$

③ SO_4^{2-}

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

$x - 8 = -2$

$x = 8 - 2$

$\boxed{x = 6}$

④ NH_3 → non metal

$x + 3(+1) = 0$

$x + 3 = 0$

$\boxed{x = -3}$

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⑤ HNO_3

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

$x + 1 - 6 = 0$

$x = 6 - 1$

$\boxed{x = 5}$ ✓

⑥ N_2O_5

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

$2x - 10 = 0$

$2x = 10$

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

$\boxed{x = 5}$ ✓

⑦ PO_4^{3-}

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

$x - 8 = -3$

$x = 8 - 3$

$\boxed{x = 5}$ ✓

⑧ NH_4^+

$x + 4(+1) = 1$

$x + 4 = 1$

$x = 1 - 4$

$\boxed{x = -3}$ ✓

Scanned with CamScanner

(9) $\underline{Cl}O_4^-$

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

$x - 8 = -1$

$x = -1 + 8$

$x = 8 - 1$

$\boxed{x = 7}$ ✓

(10) $\underline{C}H_4$

$x + 4(+1) = 0$

$\boxed{x = -4}$ ✓

(11) $\underline{N_2}O_3$

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

$2x - 6 = 0$

$2x = 6$

$\boxed{x = 3}$ ✓

(12) $H_3\underline{P}O_2$

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

$x + 3 - 4 = 0$

$x = 4 - 3$

$\boxed{x = 1}$ ✓

(13) $\underline{Mg}(OH)_2$

$x + 2(-1) = 0$

$x - 2 = 0$

$\boxed{x = 2}$ ✓

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(14) Na

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

(15) O_2

~~$2(x) \neq 0$~~

~~$x = 0$~~

~~$2(-2) = 0$~~

~~$\boxed{-4}$~~

$2x = 0$

$\boxed{x = 0}$ ✓

(16) H_2O_2

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

$2x + 2 = 0$

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

$\boxed{x = -1}$ ✓

Q Compare Acidic Strength

~~Ex~~ (1) SO_2 SO_3

$x + 2(-2) = 0$ $x = 6$

$x = 04$

$\boxed{SO_2 < SO_3}$ ✓

(2) N_2O NeO N_2O_3 NO_2 N_2O_5

1 2 3 4 5

$\boxed{N_2O < NO < N_2O_3 < NO_2 < N_2O_5}$ ✓

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③ $HClO$ $HClO_3$ $HClO_4$

1 5 7

$$HClO < HClO_3 < HClO_4$$

④ H_2SO_4 H_2SO_3

6 4

$$H_2SO_3 < H_2SO_4$$

⑤ SnO SnO_2

2 4

$$SnO < SnO_2$$

Hydra Acids

→ Acids which contains hydrogen with other non metal

eg.

EN↑; Acid strength ↑ →

Size ↑

Acid Strength ↑

NH_3	<	H_2O	<	HF
PH_3	<	H_2S	<	HCl
AsH_3	<	H_2Se	<	HBr
SbH_3	<	H_2Te	<	HI

down the group Bond strength ↓

So HI will release H^+ ion first.

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

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

$$A^+_{(g)} + B^-_{(g)} \longrightarrow AB_{(s)}$$

$\Delta H_{L.E}$ = energy released
⊖ ve

→ It is the amount of energy required to break one mole solid/latice ~~the~~ ints their constituent gasious ions.

$$AB_{(s)} \longrightarrow A^+_{(g)} + B^-_{(g)}$$

$\Delta H_{L.E}$ = energy required
⊕ ve

factors of L.E :-

$L.E \propto q_1 q_2$ { q_1 → charge on cation, q_2 → charge on anion }

$L.E \propto \frac{1}{r}$ { r ⇒ internuclear distance }

$r \simeq r_A + r_C$ → radius of cation
r_A → radius of anion

If $|q_1 q_2|$ or amount of charge ↑; L.E ↑

r ↑ ; L.E ↓

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$|q_1 q_2| \uparrow$; $r \uparrow$; L.E. $\uparrow$ { Phele charge ko dekho fir radius ko }

Q1. L.E. order ?

① $\overset{+1}{Li}\overset{-1}{H}$ $\quad$ $\overset{+1}{Na}\overset{-1}{H}$

$r \downarrow$ $\quad$ $r \uparrow$

$LiH > NaH$

② $\overset{+1}{Li}\overset{-1}{F}$ $\quad$ $\overset{+1}{Li_2}\overset{-2}{O}$ $\quad$ $\overset{+1}{Li_3}\overset{-3}{N}$

$LiF < Li_2O < Li_3N$

③ NaF $\quad$ MgF_2 $\quad$ AlF_3

$NaF < MgF_2 < AlF_3$

④ NaF $\quad$ MgO $\quad$ AlN $\quad$ SiC

$NaF < MgO < AlN < SiC$

⑤ NaF $\quad$ $NaCl$ $\quad$ $NaBr$

$NaBr < NaCl < NaF$

⑥ LiF $\quad$ NaF $\quad$ KF $\quad$ RbF $\quad$ CsF

$CsF < RbF < KF < NaF < LiF$

⑦ $\overset{+3}{Sc}\overset{-3}{N}$ $\quad$ $\overset{+4}{Ti}\overset{-4}{C}$

$ScN < TiC$

$N^{3-} \rightarrow$ Nitride

$C^{4-} \rightarrow$ Carbide

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Hydration Energy

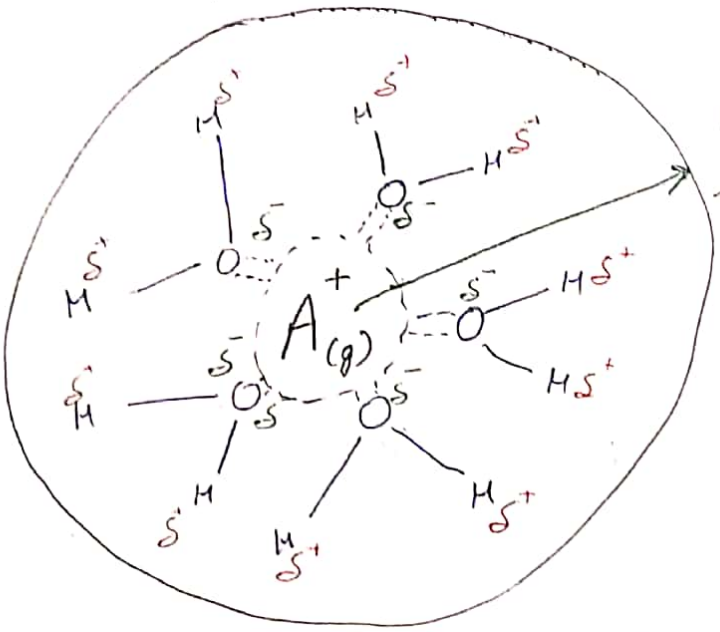
→ It is the amount of energy released when 1 mole of gaseous ions are interacted with water. (dissolve)

$$A^{+}_{(g)} + H_2O \longrightarrow A^{+}_{(aq)}$$

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

$$B^{-}_{(g)} + H_2O \longrightarrow B^{-}_{(aq)}$$

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



Hydration of Cation radius of $A^{+}_{(aq)}$

More water molecule interact ↑; ~~ions~~ hydration energy ↑.
radius of aqueous ↑

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

$$NaCl_{(s)} \xrightarrow{\Delta H_{sol}} Na^{+}_{(aq)} + Cl^{-}_{(aq)}$$

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

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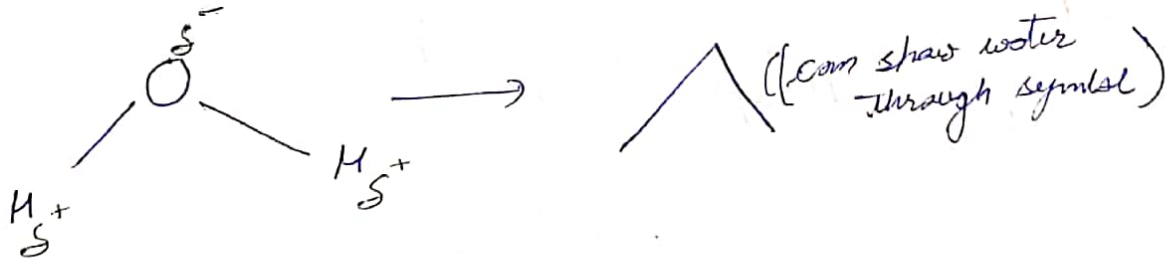
Q. $NaCl_{(s)} \longrightarrow Na^{+}_{(g)} + Cl^{-}_{(g)}$

$MgCl_{2\,(s)} \longrightarrow Mg^{2+}_{(g)} + 2Cl^{-}_{(g)}$

$AlF_{3\,(s)} \longrightarrow Al^{3+}_{(g)} + 3F^{-}_{(g)}$

$Al_2O_{3\,(s)} \longrightarrow 2Al^{3+}_{(g)} + 3O^{2-}_{(g)}$

$A_x B_{y\,(s)} \longrightarrow xA^{y+}_{(g)} + yB^{x+}_{(g)}$



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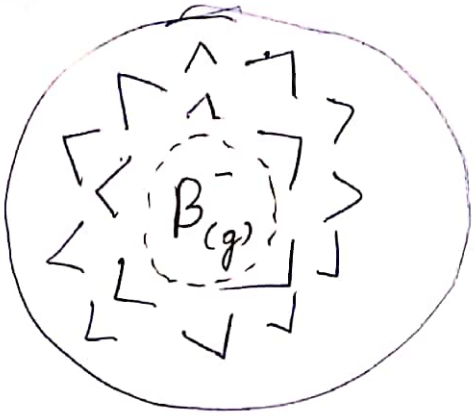

Hydration of anion

Factors affecting Hydration Energy

$H.E \propto q_1 q_2$ {charg ↑ H.E ↑}

$H.E \propto \frac{1}{r}$ {r ↑ H.E ↓}
↳ radius of ion

(charge is more dominating)

Q Compare Hydration Energy

① $Li^+_{(aq)}$ $Na^+_{(aq)}$ $K^+_{(aq)}$ $Rb^+_{(aq)}$ $Cs^+_{(aq)}$

$Li^+ > Na^+ > K^+ > Rb^+ > Cs^+$

② $F^-_{(aq)}$ $Cl^-_{(aq)}$ $Br^-_{(aq)}$ $I^-_{(aq)}$

$F^- > Cl^- > Br^- > I^-$

③ $Na^+_{(aq)}$ $Mg^{2+}_{(aq)}$ $Al^{3+}_{(aq)}$

$Na^+ < Mg^{2+} < Al^{3+}$

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(4) $F^-_{(aq)}$ $O^{2-}_{(aq)}$ $N^{3-}_{(aq)}$

$F^- < O^{2-} < N^{3-}$

Application of hydration energy :-

(1) Hydrated radius -

H.E. ↑ ; Number of water molecules ↑ Hydrated Radius ↑

eg.. Compare hydrated radius

$Na^+_{(g)} > K^+_{(g)} > Rb^+_{(g)} > Cs^+_{(g)}$ (hydrated radius)

eg (2).

$F^-_{(g)} > Cl^-_{(g)} > Br^-_{(g)} > I^-_{(g)}$ (hydrated radius)

$$\boxed{\text{Hydrated Radius} \propto H.E.}$$

(2) Ionic Mobility -

→ Ionic mobility means movement of ions.

H.E. ↑ ; more no. of H_2O interact ↑

hydrated radius ↑ ; Movement of ions ↓

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

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H.W. 11-05-2024

O-2, Mains & Advance full.

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③ Electric Conduction

Ionic mobility ∝ Electrical conductivity.

H.E ↑ Hydrated Radius ↑ Ionic mobility ↓ Electrical conductivity ↓

Conduction – Flow of Charge per unit time.

Born - Haber Cycle

Imp Terms -

① Enthalpy of Sublimation (ΔH_{sub}) - Amount of energy required/ absorbed when 1 mole of solid is converted to gas.

$$A_{(s)} \longrightarrow A_{(g)} \quad \Delta H_{sub} = \oplus ve \text{ (endothermic)}$$

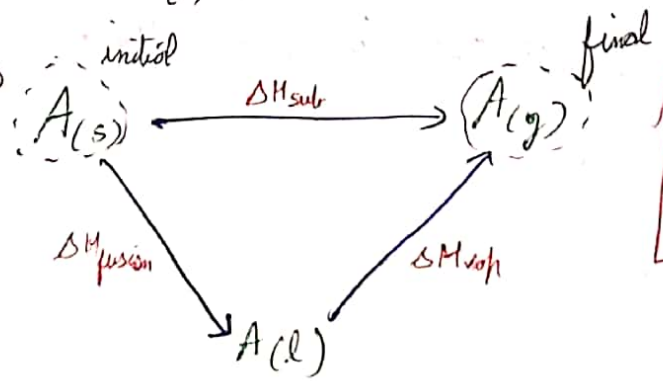
② Enthalpy of Vaporisation (ΔH_{vap}) - Amount of energy supplied/ absorbed when one mole of liquid is converted into gas.

$$A_{(l)} \longrightarrow A_{(g)} \quad \Delta H_{vap} = \oplus ve \text{ (Endothermic)}$$

③ Enthalpy of Fusion (ΔH_{fusion}) - Amount of energy required/ supplied to convert 1 mole of solid into liquid.

$$A_{(s)} \longrightarrow A_{(l)} \quad \Delta H_{fusion} = \oplus ve \text{ (exothermic)}$$

Hess law →



$$\Delta H_{sub} = \Delta H_{fusion} + \Delta H_{vap}$$

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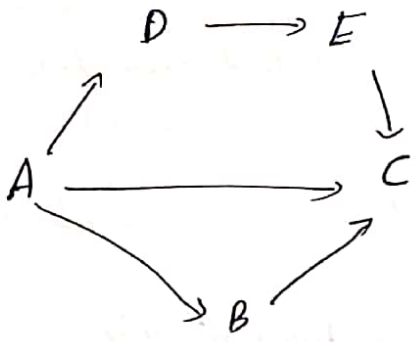
Hess Law -



$$\Delta H_{AC} = \Delta H_{ADEC} = \Delta H_{ABC}$$

$$\Delta H_{AC} = \Delta H_{AD} + \Delta H_{DE} + \Delta H_{CE}$$

$$\Delta H_{AC} = \Delta H_{AB} + \Delta H_{BC}$$

* 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 / natural state.

$$Na_{(s)} + \frac{1}{2}Cl_{2(g)} \longrightarrow NaCl_{(s)}$$

$$\Delta H_{f\,(NaCl_{(s)})} = -ve$$

Standard states -

Metal ⇒ Solid State (Exception - Hg (Mercury) → Liquid.)

O_2 ⇒
Cl_2 ⇒
N_2 ⇒
H_2 ⇒
] Gas

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Br_2 → Liquid

I_2 → Solid

Carbon → Solid (Graphite)

S_8 → Solid

P_4 → Solid

Eg. ① $\Delta H_f (MgO_{(s)})$

$$Mg_{(s)} + \frac{1}{2} O_{2(g)} \longrightarrow MgO_{(s)}$$

② $\Delta H_f (AlF_{3(s)})$

$$Al_{(s)} + \frac{3}{2} F_{2(g)} \longrightarrow AlF_{3(s)}$$

③ $\Delta H_f (MgCl_{2(s)})$

$$Mg_{(s)} + Cl_{2(g)} \longrightarrow MgCl_{2(s)}$$

④ $\Delta H_f (NaBr_{2(s)})$

$$Na_{(s)} + \frac{1}{2} Br_{2(g)/(l)} \longrightarrow NaBr_{(s)}$$

⑤ $\Delta H_f (NaI_{(s)})$

$$Na_{(s)} + \frac{1}{2} I_{2(s)} \longrightarrow NaI_{(s)}$$

⑥ $\Delta H_f (Ca_3N_{2(s)})$

$$3\,Ca_{(s)} + N_{2(g)} \longrightarrow Ca_3N_{2(s)}$$

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⑤ Bond Dissociation energy (B.D.E) – Amount of energy supplied to break one mole bond of gaseous molecules into their constituent gaseous atoms.

$$F_{2(g)} \longrightarrow 2F_{(g)} \quad \Delta H_{B.D.E(F_2(g))}$$

$$\frac{1}{2}Cl_{2(g)} \longrightarrow Cl_{(g)} \quad \frac{1}{2}\Delta H_{B.D.E(Cl_2(g))}$$

$$Br_{2(l)} \longrightarrow Br_{(l)}$$

$$\Delta H_{Reaction} \neq \Delta H_{B.D.E(Br_2(l))}$$

$$Br_{2(g)} \rightarrow 2Br_{(g)}$$

$$\Delta H_{reaction} = \Delta H_{B.D.E(Br_2(g))}$$

Born Haber Energy of $NaCl_{(s)}$

$$Na_{(s)} + \frac{1}{2}Cl_{2(g)} \longrightarrow NaCl_{(s)} \quad \Delta H_{f(NaCl(s))}$$

$Na_{(s)} \xrightarrow{\Delta H_{sub}\ (+)} Na_{(g)}$

$\frac{1}{2}Cl_{2(g)} \xrightarrow{\frac{1}{2}\Delta H_{BDE(Cl_2(g))}\ (+)} Cl_{(g)}$

$Na_{(g)} \xrightarrow{\Delta H_{I.E.}\ (+)} Na^+_{(g)}$

$Cl_{(g)} \xrightarrow{\Delta H_{EGD}\ (-)} Cl^-_{(g)}$

$$Na^+_{(g)} + Cl^-_{(g)} \xrightarrow{\Delta H_{L.E}} NaCl_{(s)}$$

$$\Delta H_{f(NaCl(s))} = \Delta H_{sub(Na_{(s)})} + \Delta H_{I.E(Na_{(g)})} + \frac{1}{2}\Delta H_{B.D.E(Cl_2(g))} + \Delta H_{EGD(Cl_{(g)})} + \Delta H^{\ominus}_{L.E(NaCl_{(s)})}$$

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② write Born-Haber cycle of $MgO_{(s)}$

$$Mg_{(s)} + \frac{1}{2} O_{2(g)} \longrightarrow MgO_{(s)}$$

$$Mg_{(s)} \xrightarrow{\Delta H_{sub}\ (+ve)} Mg_{(g)}$$

$$\frac{1}{2} O_{2(g)} \xrightarrow{\frac{1}{2}\Delta H_{B.D.E}(O_{2(g)})\ (+ve)} O_{(g)}$$

$$Mg_{(g)} \xrightarrow{\Delta H_{I.E}\ (+ve)} Mg^{+}_{(g)}$$

$$O_{(g)} \xrightarrow{\Delta H_{EGE}\ (-ve)} O^{-}_{(g)}$$

$$Mg^{+}_{(g)} \xrightarrow{\Delta H_{I.E_2}\ (+ve)} Mg^{2+}_{(g)}$$

$$O^{-}_{(g)} \xrightarrow{\Delta H_{EGE_2}\ (+ve)} O^{2-}_{(g)}$$

$$Mg^{2+}_{(g)} + O^{2-}_{(g)} \xrightarrow{\Delta H_{L.E.}\ (-ve)} MgO_{(s)}$$

$$\Delta H_f(MgO_{(s)}) = \Delta H_{sub}(Mg_{(s)}) + \Delta H_{I.E.}(Mg_{(g)}) + \Delta H_{I.E.}(Mg^{+}_{(g)})$$

$$+ \frac{1}{2}\Delta H_{B.D.E}(O_{2(g)}) + \Delta H_{EGE}(O_{(g)}) + \Delta H_{EGE}(O^{-}_{(g)})$$

$$+ \Delta H_{L.E.}$$

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Q Born Haber cycle for $AlF_{3(s)}$

$$Al_{(s)} + \frac{3}{2} F_{2(g)} \longrightarrow AlF_{3(s)}$$

$\downarrow \Delta H_{sub}$ (+) $\qquad \downarrow \frac{3}{2}\Delta H_{B.D.E}(F_{2(g)})$ (+)

$Al_{(g)} \qquad 3F_{(g)}$

$\downarrow \Delta H_{I.E_1}$ (+) $\qquad \downarrow 3\Delta H_{EGE}$ (−)

$Al^{+}_{(g)} \qquad 3F^{-}_{(g)}$

$\downarrow \Delta H_{I.E_2}$ (+)

$Al^{2+}_{(g)}$

$\downarrow \Delta H_{I.E_3}$ (+)

$Al^{3+}_{(g)} + 3F^{-}_{(g)} \xrightarrow[\Delta H_{L.E.}]{} AlF_{3(s)}$ (−)

Q (4) $MgCl_{2(s)}$ H.W. 13-05-2024

Roce - till 14

(5) $AlN_{(s)}$

S-2

(6) $Ca_3N_{2(s)}$

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Solublity of Ionic compound in (H_2O)

$$NaCl_{(s)} \longrightarrow Na^+_{(aq)} + Cl^-_{(aq)}$$

$$NaCl_{(s)} \xrightarrow{\Delta H_{L.E}} Na^+_{(g)} + Cl^-_{(g)}$$

$$Na^+_{(g)} \xrightarrow{\Delta H_{H.E}} Na^+_{(aq)}$$

$$Cl^-_{(g)} \xrightarrow{\Delta H_{H.E.}} Cl^-_{(aq)}$$

$\boxed{|\Delta H_{L.E}| > |\Delta H_{H.E}|}$ Not Soluble

$\boxed{|\Delta H_{L.E}| < |\Delta H_{H.E}|}$ soluble in H_2O

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H.W.

① $MgCl_2$

$$Mg_{(s)} + Cl_{2(g)} \longrightarrow MgCl_{2(s)}$$

$Mg_{(s)} \xrightarrow{\Delta H_{sub.} \ (+)} Mg_{(g)}$

$Cl_{2(g)} \xrightarrow{\Delta H_{B.D.E} \ (+)} 2Cl_{(g)}$

$Mg_{(g)} \xrightarrow{\Delta H_{I.E} \ (+)} Mg^{+}_{(g)}$

$2Cl_{(g)} \xrightarrow{2\Delta H_{EGE} \ (-)} 2Cl^{-}_{(g)}$

$Mg^{+}_{(g)} \xrightarrow{\Delta H_{I.E_2} \ (+)} Mg^{2+}_{(g)}$

$$Mg^{2+}_{(g)} + 2Cl^{-}_{(g)} \xrightarrow{\Delta H_{L.E} \ (-)} MgCl_{2(s)}$$

② $AlN_{(s)}$

$$Al_{(s)} + \frac{1}{2}N_{2(g)} \longrightarrow AlN_{(s)}$$

$Al_{(s)} \xrightarrow{\Delta H_{sub} \ (+)} Al_{(g)}$

$\frac{1}{2}N_{2(g)} \xrightarrow{\Delta H_{BDE} \ (+)} N_{(g)}$

$N_{(g)} \xrightarrow{\Delta H_{EGE} \ (+)} N^{-}_{(g)}$

$Al_{(g)} \xrightarrow{\Delta H_{I.E} \ (+)} Al^{+}_{(g)}$

$Al^{+}_{(g)} \xrightarrow{\Delta H_{I.E_2} \ (+)} Al^{2+}_{(g)}$

$N^{-}_{(g)} \xrightarrow{\Delta H_{EGE_2} \ (+)} N^{2-}_{(g)}$

$Al^{2+}_{(g)} \xrightarrow{\Delta H_{I.E_3} \ (+)} Al^{3+}_{(g)}$

$N^{2-}_{(g)} \xrightarrow{\Delta H_{EGE_3} \ (+)} N^{3-}_{(g)}$

$$Al^{3+}_{(g)} + N^{3-}_{(g)} \xrightarrow{\Delta H_{L.E.} \ (-)} AlN_{(s)}$$

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③ Ca_3N_2

$$3Ca_{(s)} + N_{2(g)} \longrightarrow Ca_3N_{2(s)}$$

$$\downarrow 3\Delta H_{sub} \oplus \qquad \downarrow \Delta H_{B.D.E} \oplus$$

$$3Ca_{(g)} \qquad 2N_{(g)}$$

$$\downarrow 3\Delta H_{I.E} \oplus \qquad \downarrow 2\Delta H_{E.G.E} \oplus$$

$$3Ca^{+}_{(g)} \qquad 2N^{-}_{(g)}$$

$$\downarrow 3\Delta H_{I.E_2} \oplus \qquad \downarrow 2\Delta H_{E.G.E_2} \oplus$$

$$2N^{2-}_{(g)}$$

$$\downarrow 2\Delta H_{E.G.E_3} \oplus$$

$$3Ca^{2+}_{(g)} + 2N^{3-}_{(g)} \xrightarrow[\Delta H_{L.E.} \ominus]{} Ca_3N_{2(s)}$$

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Chemical Bonding

Chemical bond - It is the force of attraction between two atoms which binds them together.

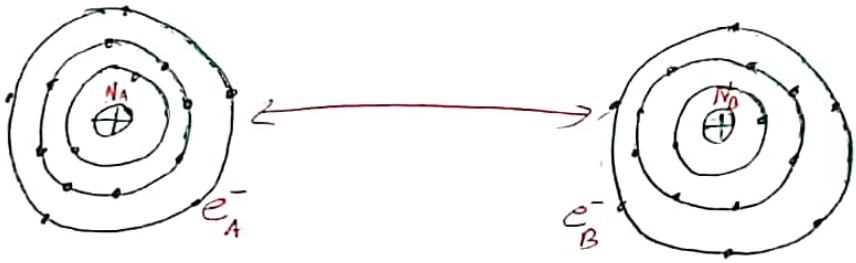
Why a chemical bond is formed?

→ To gain stability ~~to~~ ~~due~~ by decreasing their potential energy or to aquire inert gas configuration.

→ Atoms are more reactive than molecules because they have higher energy than molecules.

Cl → atom (more reactive)

Cl_2 → molecule (less reactive)



Attraction	Repulsion
$N_A - e_B$	$e_A - e_B$
$N_B - e_A$	$N_A - N_B$

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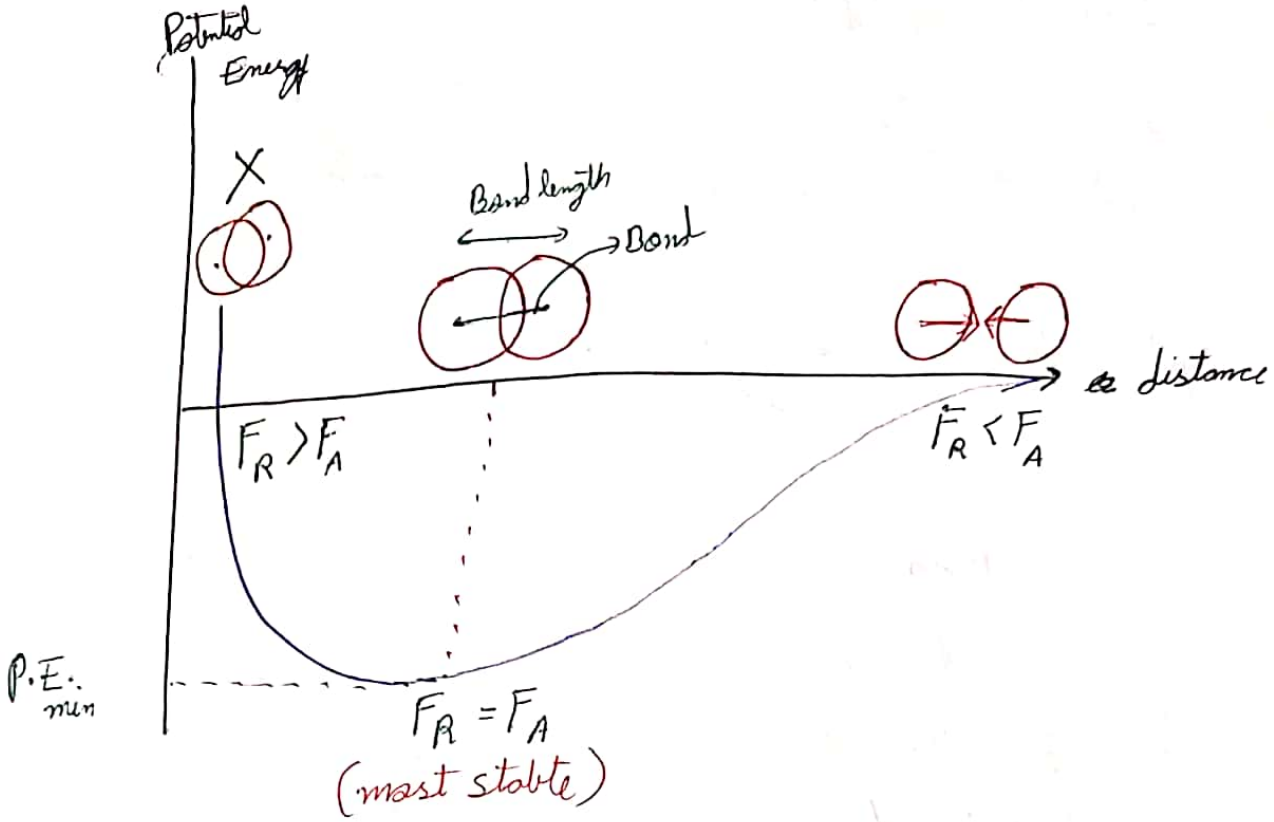

F_R = Force of Repulsion

F_A = Force of Attraction.

→ 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 amount of energy required to break bond.

or

→ It is the amount of energy released during bond formation.

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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 ($E_{Released} > 42\ kJ/mol$)

1. Ionic / Electrovalent Bond
2. Covalent Bond
3. Co-ordinate Bond / Dative Bond
4. Metallic Bond.

Weak Bond ($E_{Released} < 42\ kJ/mol$)

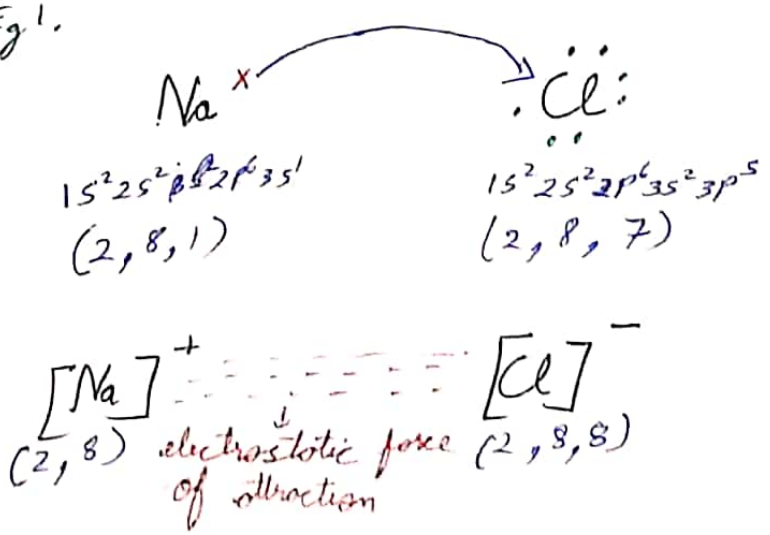
- Hydrogen Bond ($8 < E_{Rel} < 42$)
- Weak forces ($E_{Rel} < 8\ kJ/mol$)

---

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 1.



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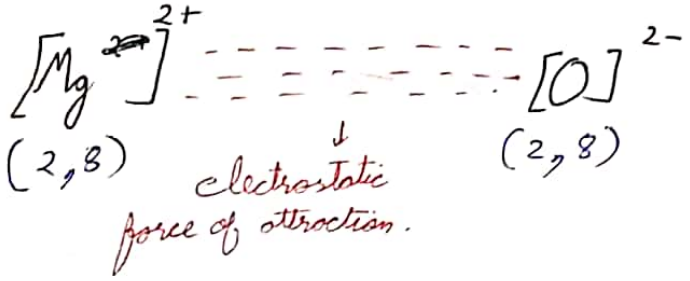
Eg 2.



→ Ionic bond is non - directional in Nature.

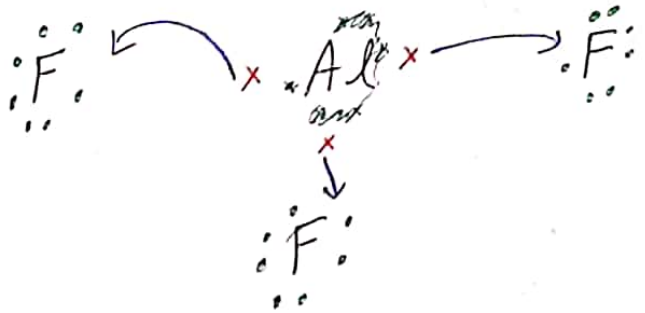
<u>Electro valency</u> → It is defined only for ionic bonds.

→ 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

2. MgO E.V (O) = 2
E.V. (Mg) = 2

3. AlF_3 E.V. (Al) = 3
E.V. (F) = 1



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Condition for the formation of ionic bond →

→ $\boxed{\Delta E.N \geq 2.1}$

→ Low ionisation energy of electropositive element.

→ High Electron Afinity of electronegative element.

→ High Lattice energy.

---

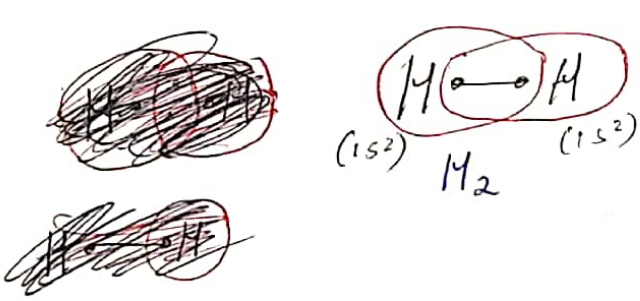
2. Covalent Bond :-

→ This type of Bond is formed by equal sharing of electrons between two bonded atoms.

~~→ Covalent Bonds are dir~~

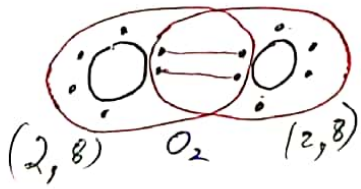
→ Covalent Bonds are directional in nature.

→ Eg 1.



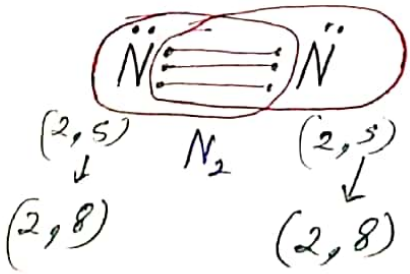
⇒ Single Bond

Eg 2.



⇒ Double Bond

Eg 3.



⇒ Triple Bond

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Electron Pair

→ Pair of e^- is known as electron pair

Electron Pair

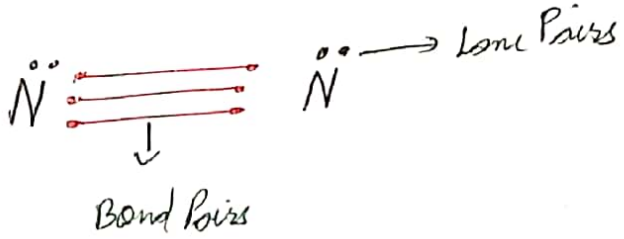
Bond Pair (B.P)

→ Pairs of e^- which are participating during bond formation

Lone Pair (L.P)

→ Pairs of e^- which do not participating during bond formation.

Eg1.



Bond Pairs - 3

Lone Pairs - 2

Q Count Bond Pairs & Lone Pairs in

1. Cl_2
2. H_2O
3. CS_2
4. SF_2
5. HCl
6. NCl_3
7. C_2F_2
8. CO_2
9. NH_3
10. CH_4

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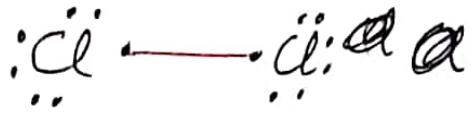
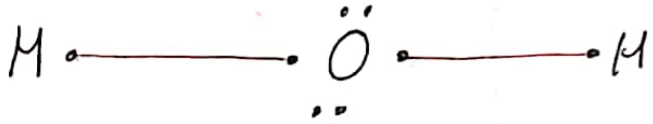
(1)



Bond Pair → 1

Lone Pair → 6

(2)

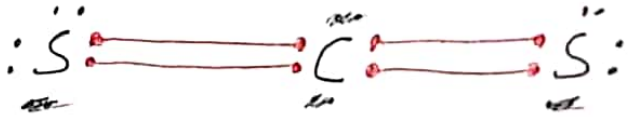


BP - 2
LP - 2

(3)

BP - 1
LP - 3

(4)



BP - 4
LP - 4

(5)

BP - 2
LP - 8

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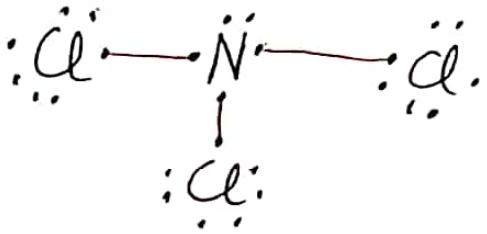
(6)



BP- 3
LP- 10

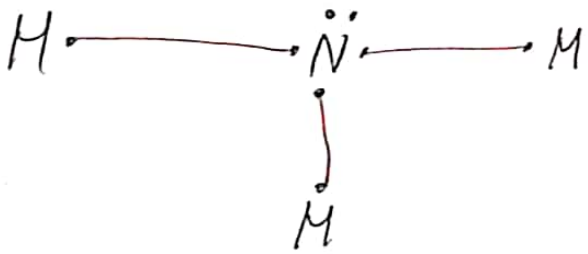
(7)

BP- 5
LP- 6

(8)

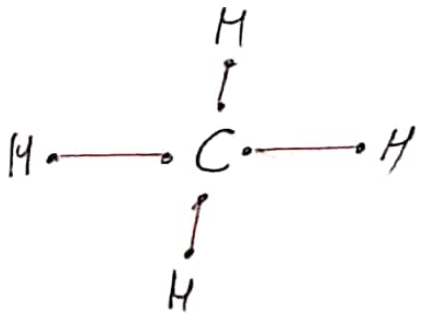
BP - 4
LP- 4

(9)



BP - 3
LP- 1

(10)



BP - 4
LP- 0

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Covalency - The number of electrons that are participating of an element in 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.

or

Number of unpaired electrons in outermost shell in ground state or in excited state decides co-valency of an atom.

or

→ 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

B : [He] $2s^2 2p^1$

[He] [↑↓] [↑ | |]

} covalency = 1 (do not exist)

B^* : [He] [↑↓] [↑ | ↑ |] C.V. = 3

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.

eg ① BCl_3, BF_3, $AlCl_3$ (covalent, not ionic bond)

$\Delta EN < 2.1$

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Carbon family - (tetravalent)

C: [He] $2s^2 2p^2$

[He] [↑↓] [↑|↑|]

C*: [↑↓] [↑|↑|↑|] covalency - 4

Eg①. CH_4, CO_2, $COCl_2$, CCl_4

→ C, Si, Ge, Sn, Pb prefer to form 4 bonds (C.V.) ⇒ 4

Nitrogen Family

N: [He] $2s^2 2p^3$, cannot excite as 2d do not exist.

[↑↓] [↑|↑|↑] - 3 unpaired (C.V. = 3)

P: [Ne] $3s^2 3p^3$

[↑↓] [↑|↑|↑] (C.V = 3 exists)

P*: [↑] [↑|↑|↑] [↑| | | | |] (C.V. = 5)
3d

~~→ Nitrogen forms a ong~~

→ Nitrogen has covalency 3.

→ P, As, Sb has covalency both 3 & 5.

Eg① PCl_3, PCl_5

Oxygen family

O: [↑↓] [↑↓|↑|↑] (C.V. = 2)

S: [↑↓] [↑↓|↑|↑] [| | | | |] (C.V = 2)
3s 3p 3d

S*: [↑↓] [↑|↑|↑] [↑| | | | |] (C.V = 4)

S**: [↑] [↑|↑|↑] [↑|↑| | | |] (C.V = 6)

O: 2
S, Se, Te : 2, 4, 6

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Halogen Family

F : [He] [↑↓] [↑↓|↑↓|↑] (C.V. = 1)

Cl : [Ne] [↑↓] [↑↓|↑↓|↑] [| | |] (C.V. = 1)

Cl* : [Ne] [↑↓] [↑↓|↑|↑] [↑| | |] (C.V. = 3)

Cl** : [Ne] [↑↓] [↑|↑|↑] [↑|↑| |] (C.V = 5)

Cl*** : [Ne] [↑] [↑|↑|↑] [↑|↑|↑|] (C.V = 7)

F : 1
Cl, Br, I : 1, 3, 5, 7

Inert Gases

Xe : [Kr] $5s^2\ 4d^{10}\ 5p^6\ 5d^0$

[↑↓] [↑↓|↑↓|↑↓|↑↓|↑↓] [↑↓|↑↓|↑↓]

Xe* : [↑↓] [↑↓|↑↓|↑] [↑|↑| |] (CV = 1)

Xe** : [↑↓] [↑↓|↑|↑] [↑|↑| |] (CV = 3)

Xe*** : [↑↓] [↑|↑|↑] [↑|↑|↑|] (CV = 5)

Xe**** : [↑↓] [↑|↑|↑] [↑|↑|↑|↑] (CV = 8)

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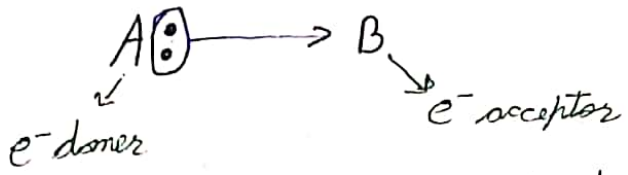
3. Co-ordinate Bond (~~Partial~~ Dative Bond)

→ 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.



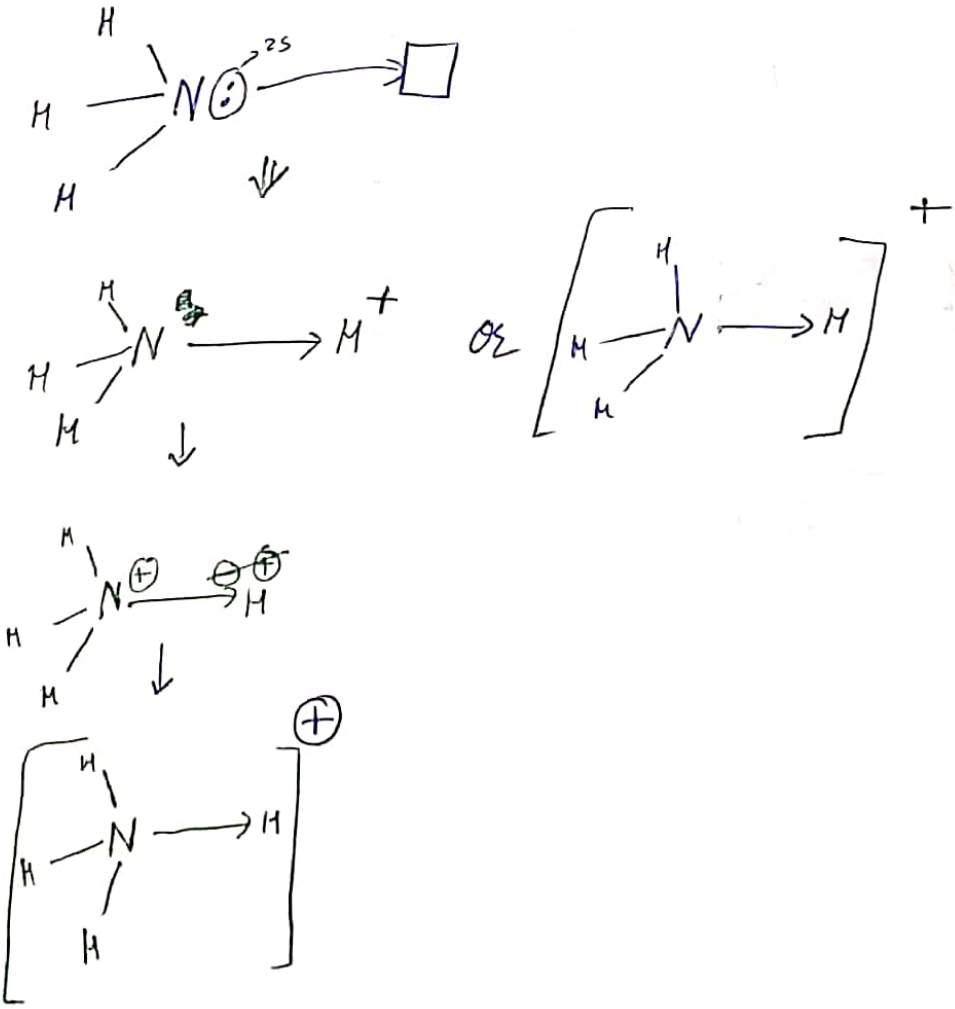
Condition to form co-ordinate bond :-

① One atom have at least 1 lone pair.

② One atom have atleast one vacant ~~empty~~ orbital.

Eg NH_4^+ (Ammonium ion)

$NH_3 + H^+$



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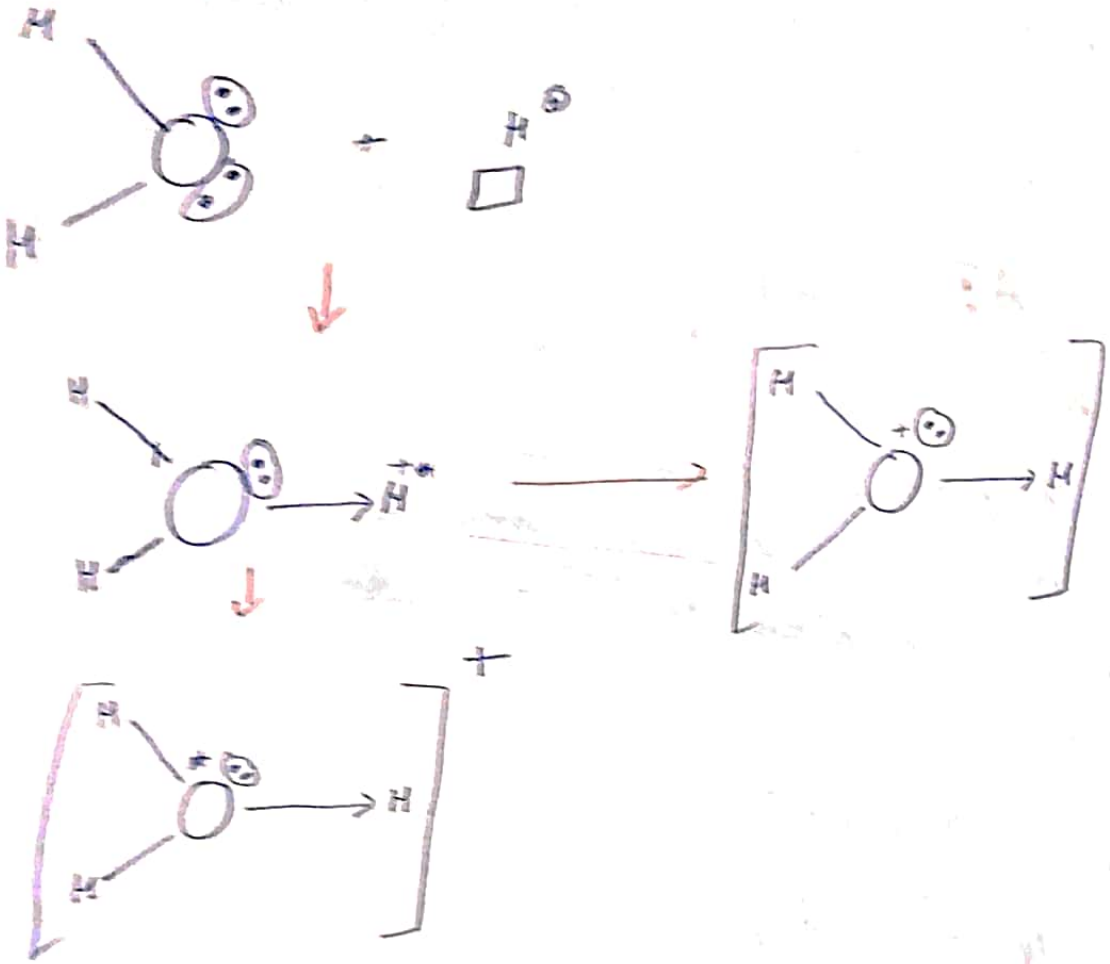
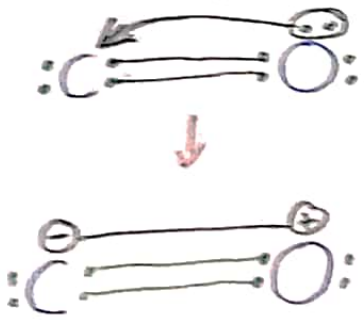
Eg ② (H_3O^+)

$H_2O + H^+$



Eg ③ CO



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Lewis Dot Structure (L-D-S):-

→ According to Lewis Theory, every atom tries to attain 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 configration (except hydrogen - Duplet)

Eg ① HF

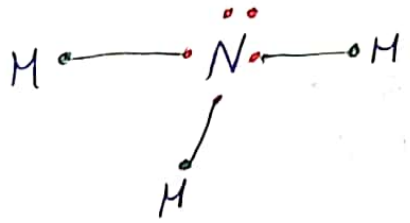
Rules to draw LDS

1. Chose ~~centre~~ Central atom -
   → Hydrogen & Flourine can never be central ~~atom~~ atom
   → Usually central atom is least electron negative atom and also lesser in number

① CO_2

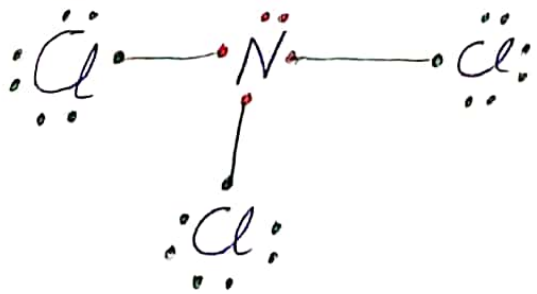
L.P. = 4
BP = 4

② NH_3



L.P = 1
B.P = 3

③ NCl_3



L.P = 1
B.P = 3

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(4) HCN (Cynite)

(5) PH_3

(6) HNC (Iso - Cynite)

(7) H_2S

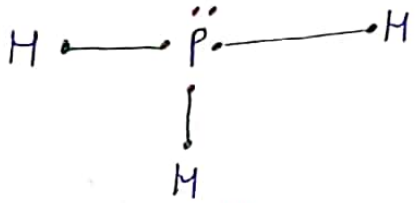
(8) SO_2

(9) SO_3

(4)

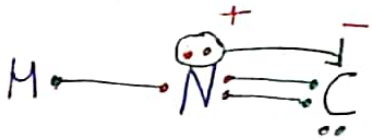
L·P - 0
B·P - 4

(5)



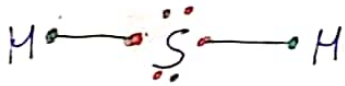
L·P - 1
B·P - 3

(6)



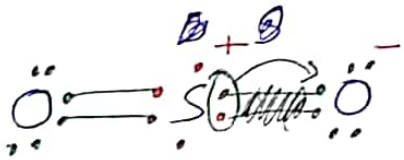
L·P - 1
B·P - 4

(7)



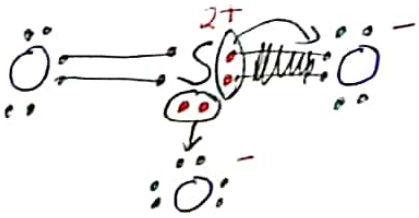
L·P - 2
B·P - 2

(8)



L·P - 6
B·P - 3

(9)



L·P - 8
B·P - 4

O: [↑↓] [↑↓] [↑↓ | ↑ | ↑]

O: [↑↓] [↑↓] [↑↓ | ↑↓ |] → To create vacant orbital for co-ordinate bond.

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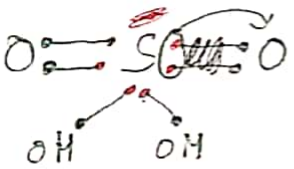
Note - ① Oxygen atom in ground state cannot form coordinate bond because it has no vacant orbital.
② However if $1e^-$ of 2p subshell gets paired in another orbital of 2p subshell then a vacant orbital is created.

(10) H_2SO_4
(11) $NOCl_3$
(12) CH_4
(13) O_3
(14) $SOCl_2$
(15) C_2F_4
(16) C_2F_2
(17) N_2O_5

(10)



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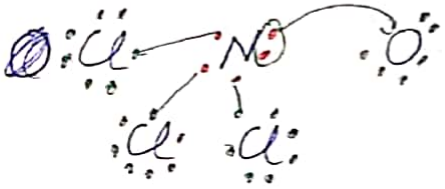

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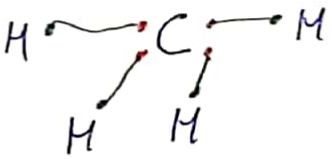

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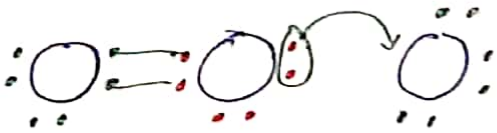

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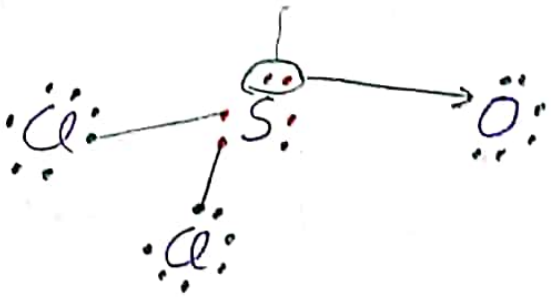

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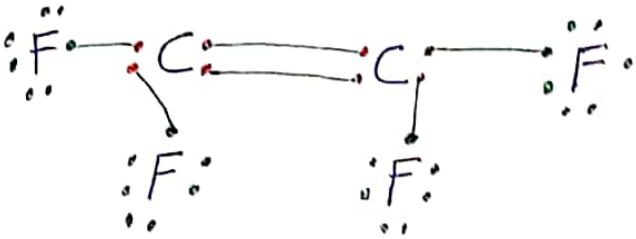
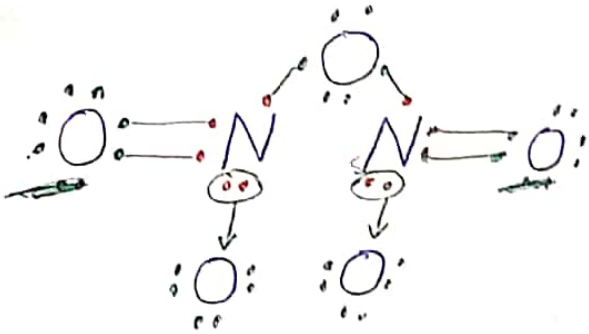


(16)

(17)



LDS of charged species

→ ⊖ve charge carried by more EN atom

→ ⊕ve charge carried by less EN atom.

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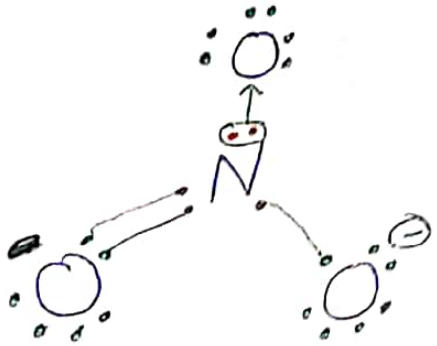
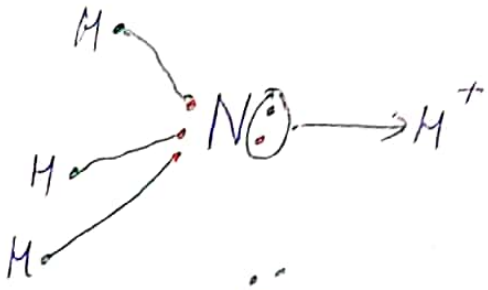
Eg ①. NO_2^-

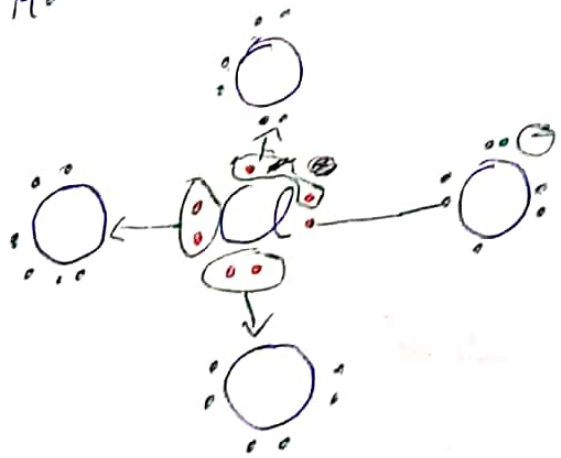
② NO_3^-



③. NH_4^+



④ ClO_4^-



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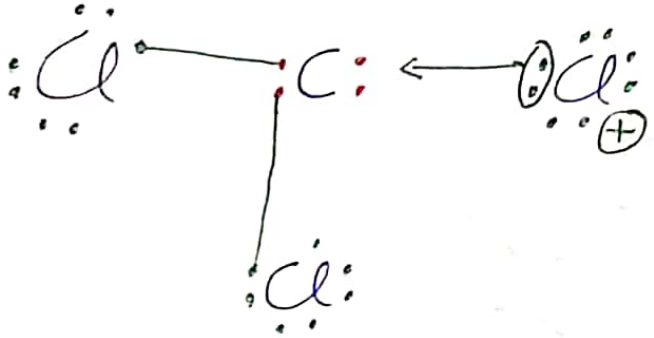
Q (5) CCl_3^-
(6) $COCl_2$
(7) N_3^-
(8) CH_3Cl
(9) NH_2Cl

(10) OCN^-
(11) CN^-
(12) SCN^-
(13) SiF_4
(14) $SnCl_3^-$

(15) BF_4^-
(16) BH_4^-
(17) BeF_4^{2-}
(18) CO_3^{2-}
(19) NO_3Cl

(20) NOCl
(Tilden's reagent)

(5)



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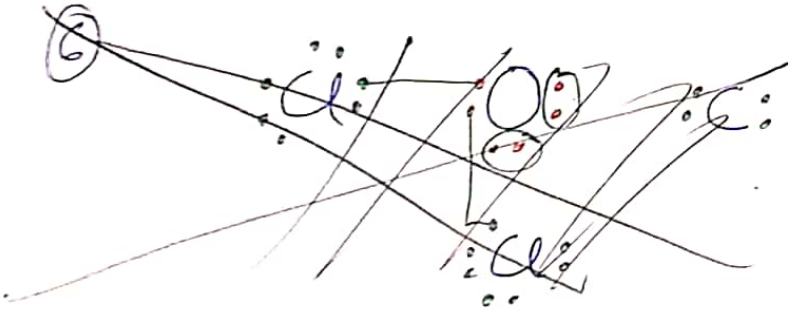

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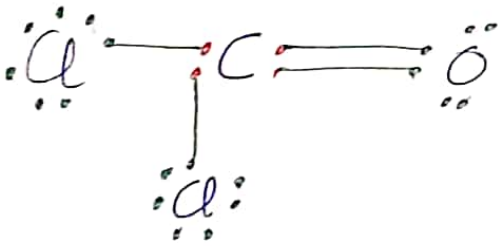

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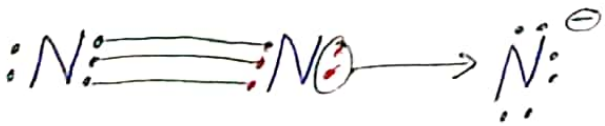

or

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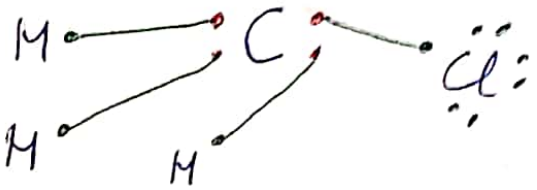

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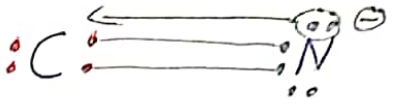

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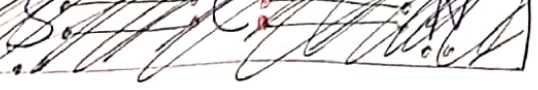

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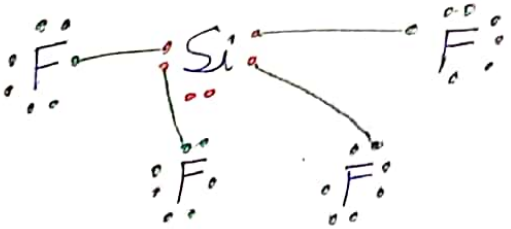

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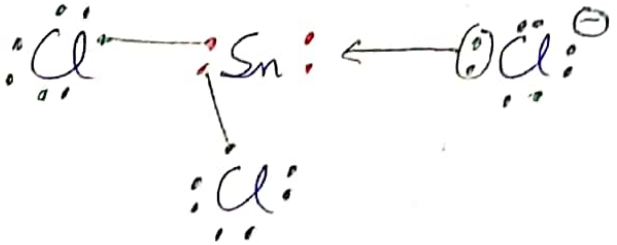

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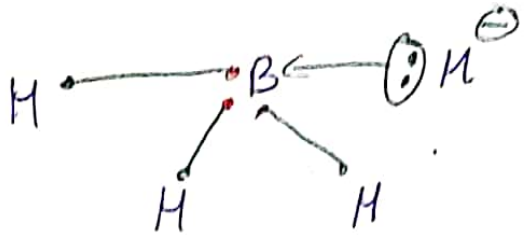

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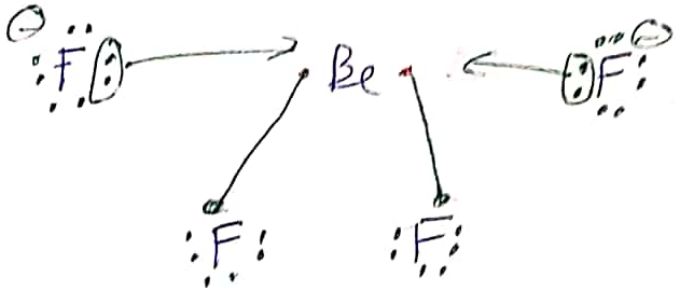

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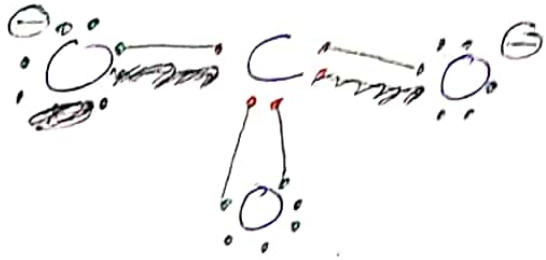

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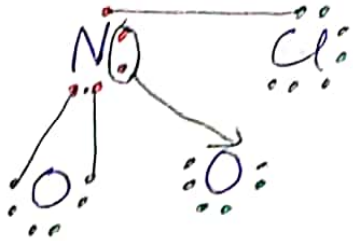

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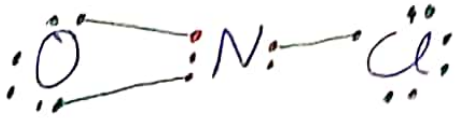

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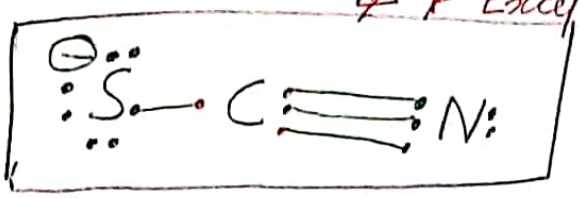
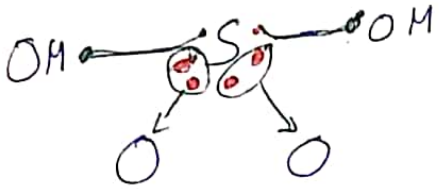
Exception



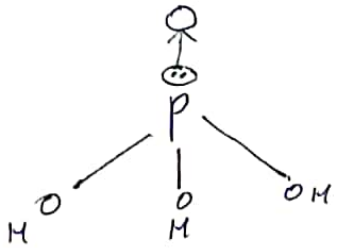
Note:-

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 covered in oxy acids).

eg ①. H_2SO_4



eg ②. H_3PO_4



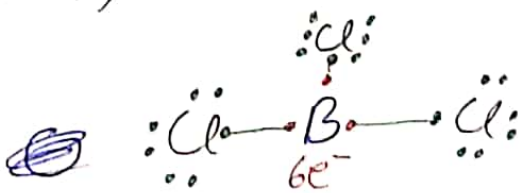
Drawbacks of Lewis Dot Structure

① Some compounds were discovered which were not following octet.

→ There are 3 categories

① hypovalent - Species in which central atom occupy less than 8 valent electrons after bonding.

eg. BF_3, BCl_3, BBr_3, BI_3, $AlCl_3$



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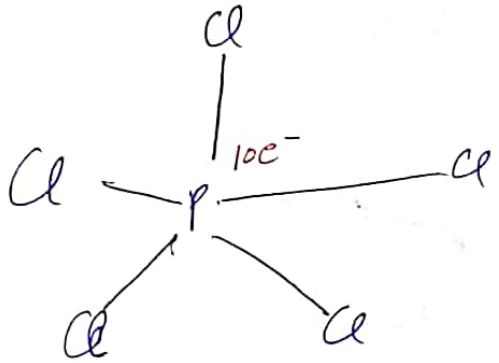
→ AlF_3 is not hypovalent because it has complete octet due to ionic nature.

→ AlF_3 is ionic but $AlCl_3$, $AlBr_3$ & AlI_3 are covalent compound. (ΔEN)

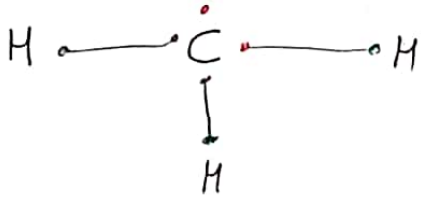
② Hyper Valent – Central atom occupies more than 8 valence e^- after bonding.

eg. PCl_5, XeF_2, XeF_4, XeF_6, IF_7, IF_5



③ odd e^- species – In which odd no. of valence e^- are present after bonding.

eg. NO, CH_3, ClO_2, NO_2, CF_3, ClO_3



② Lewis theory cannot explain shape or geometry of molecule.

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