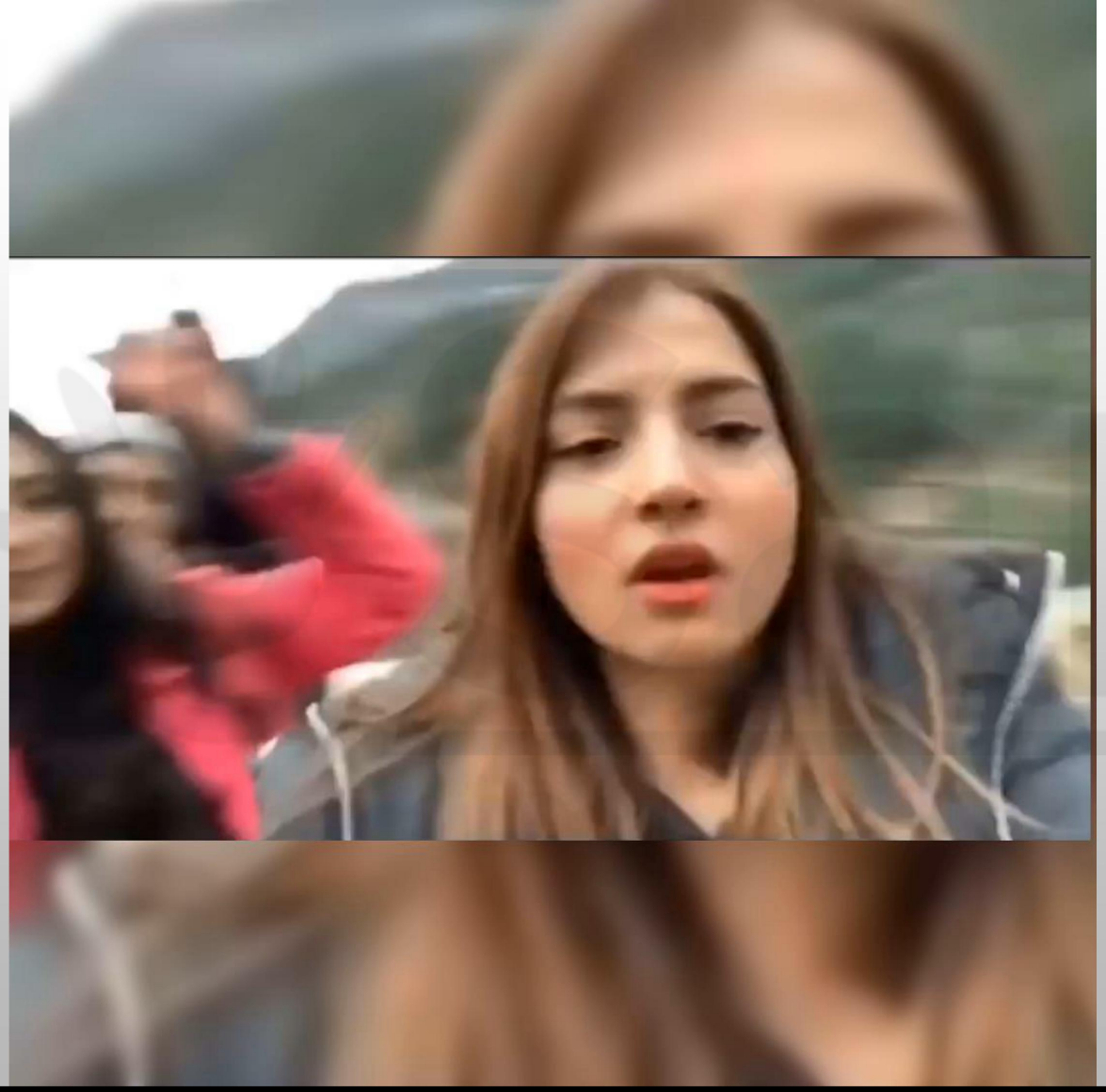

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एक Single KEY POINT

$$\frac{KQ_1Q_2}{r^2}$$

Electrostatic

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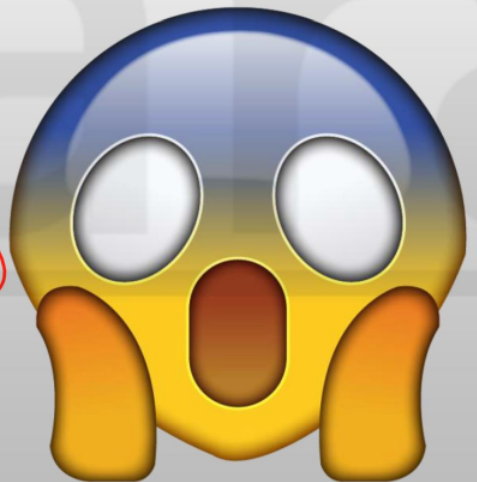
Phodng

Periodic Table → JEE/NEET X

एक सूचना जनहित में जारी

stable

There are no exceptions!!



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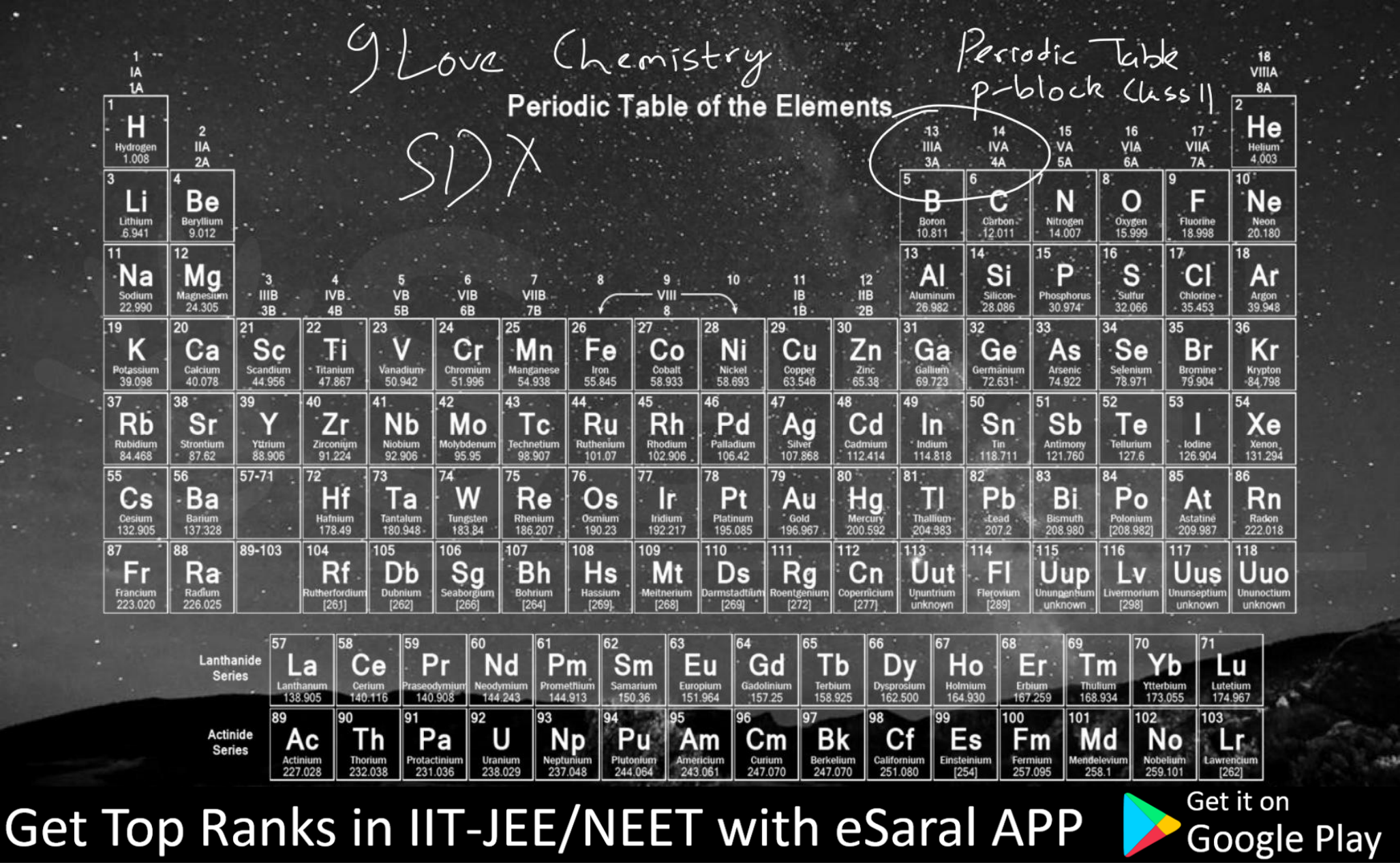
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Periodic Table of the Elements
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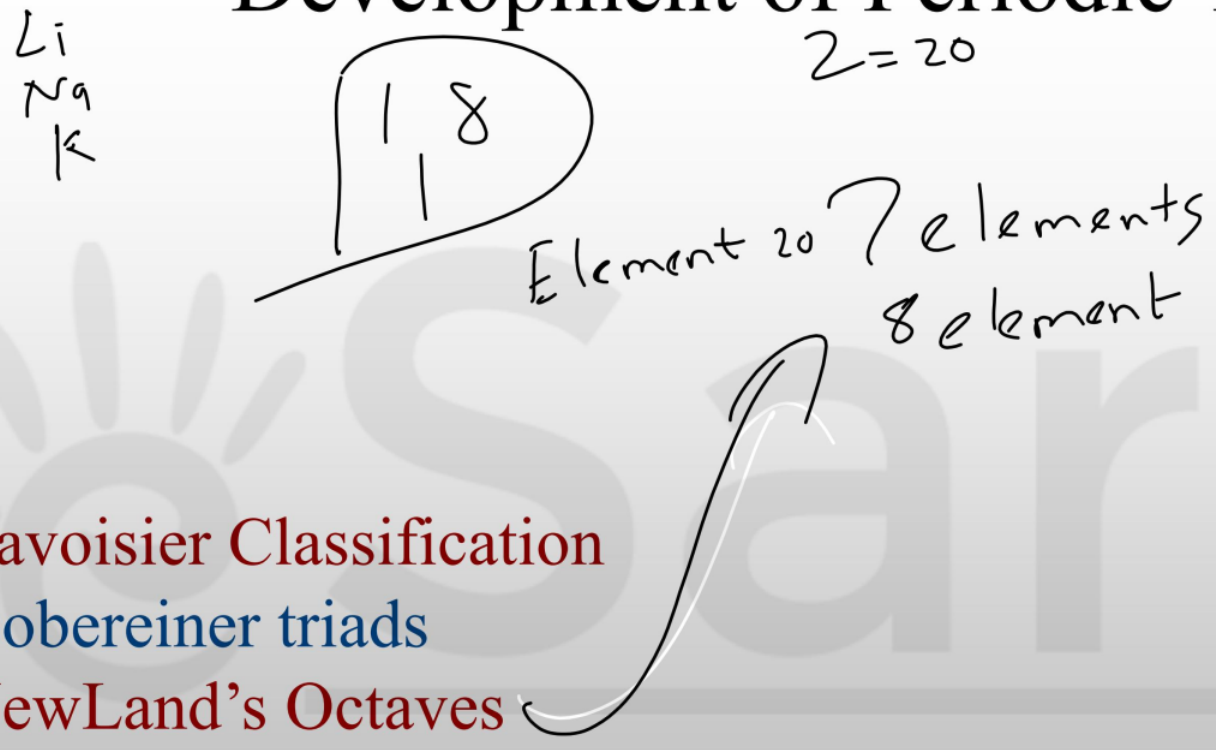
Development of Periodic Table



1. Lavoisier Classification
2. Dobereiner triads
3. NewLand's Octaves
4. Lothar Meyer Curves

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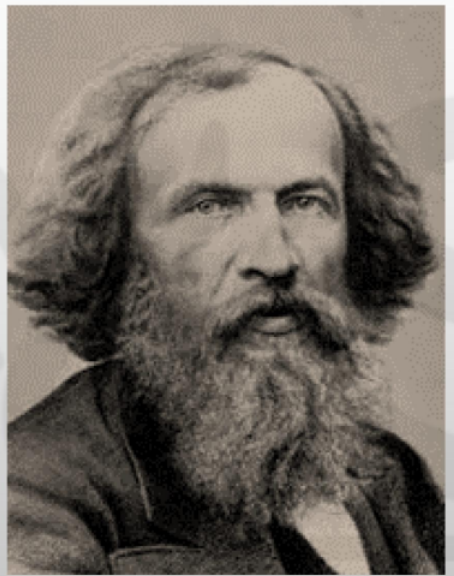

Dmitri Ivanovich Mendeleev

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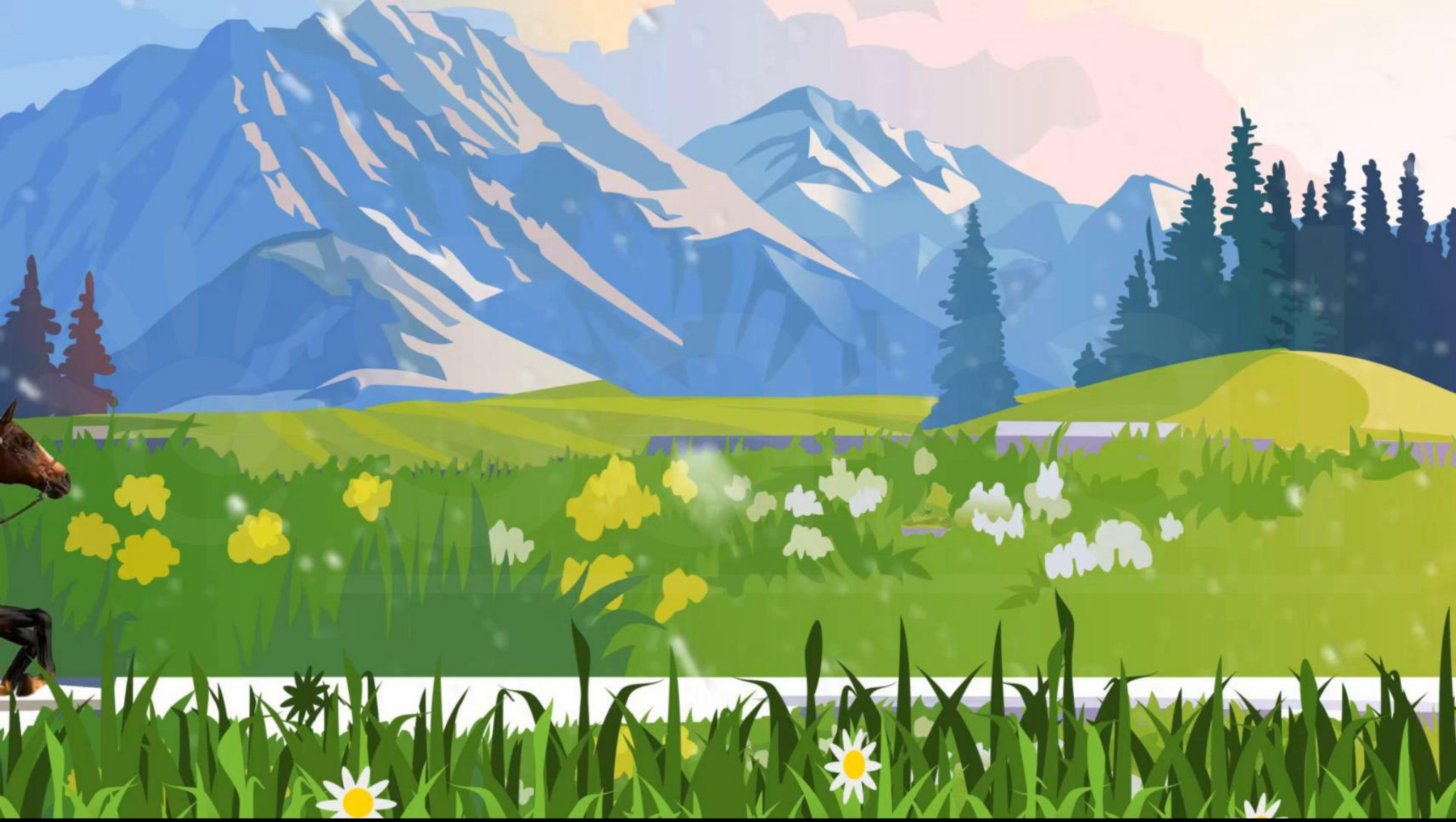

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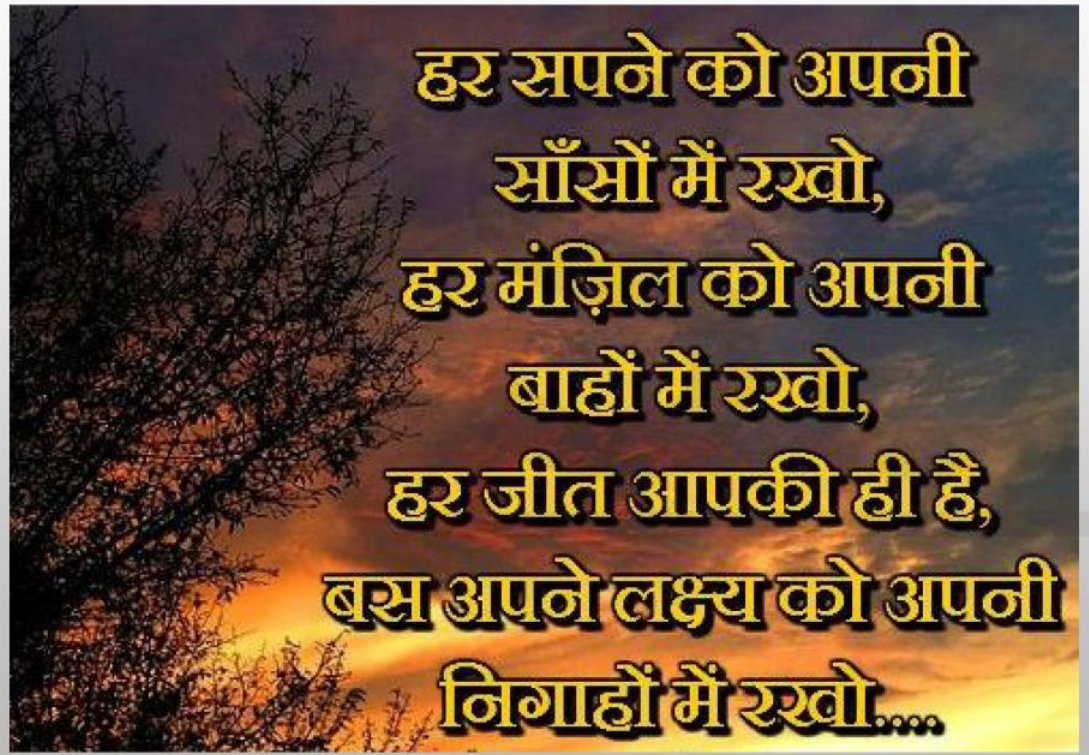

Phodna hai

↓

focus

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Characteristics of Mendeleev's periodic table :

Newland

(a) It was based on atomic weight.

(b) 63 elements were ~~known~~, noble gases were not discovered.

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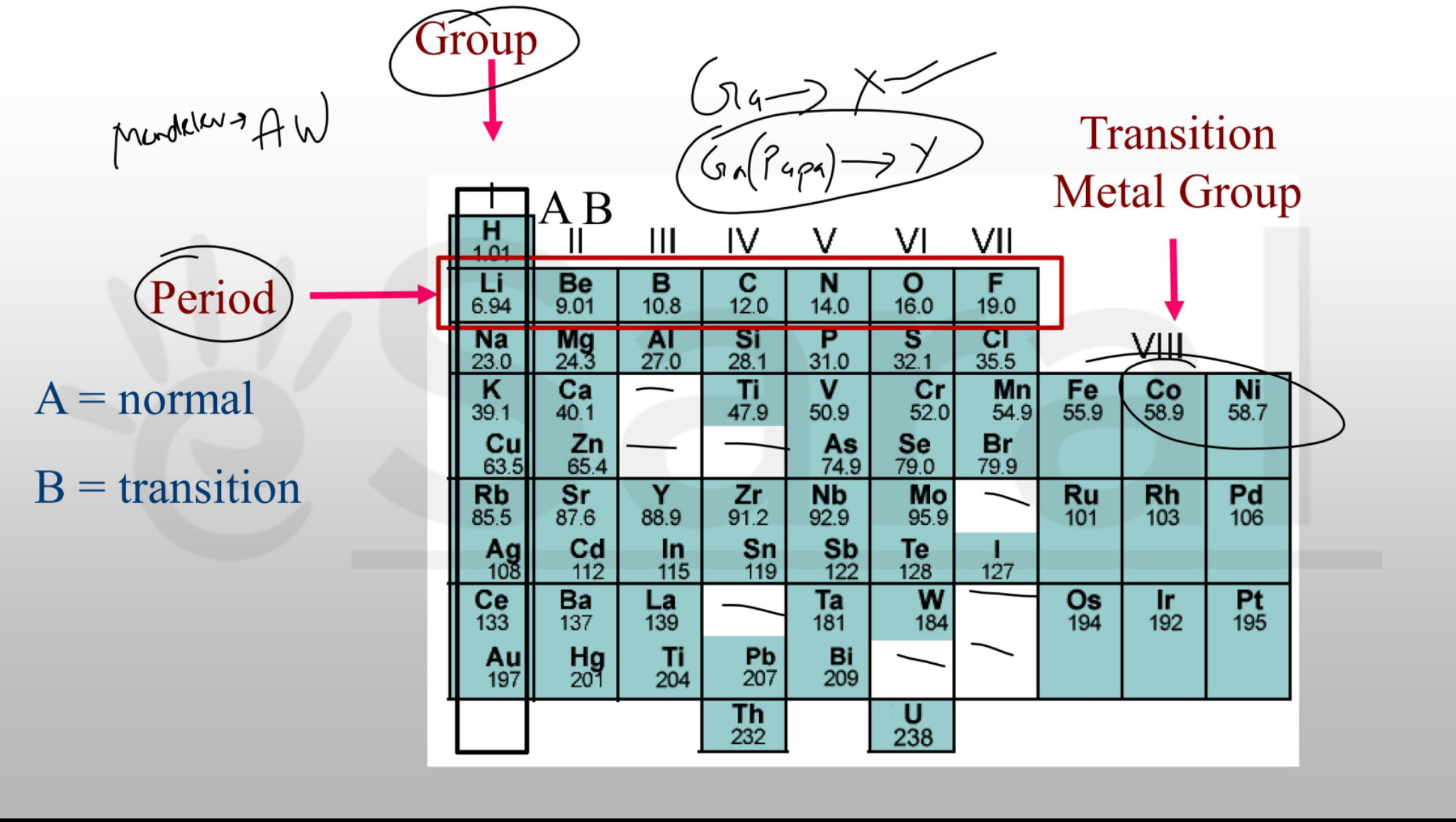

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Moseley's Discovery

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$\sqrt{\nu} \propto Z$ → Atomic Number

Qs
3rd period

Moseley found that the physical and chemical properties of the elements are periodic function of their atomic number.

This is also known as 'Modern Periodic Law'.

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Nomenclature of elements

VVImp

0	Nil
1	Un
2	Bi
3	Tri
4	Quad
5	Pent
6	Hex
7	Sept
8	Oct
9	Enn

101	unnilunium	Unu
102	unnilbium	Unb
103	unniltrium	Unt

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s-Block

Their general electronic configuration is ns^{1-2} where n = (1 to 7)

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

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

(p- Block)

(a) The general electronic configuration of p-block elements is ns^2, np^{1-6} (where n = 2 to 6)

(d- Block)

(a) The general electronic configuration of these elements is $(n-1)d^{1-10}\ ns^{1-2}$ [except Pd-$4d^{10}$, $5s^0$]where n = 4 to 7.

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f-BLOCK ELEMENTS

(a) The elements in which the last electron enters into (n–2)f-orbital are called f-block elements.

(b) The general electronic configuration of these elements is $(n-2)\ s^2\ p^6\ d^{10}\ f^{(1\text{-}14)}\ (n-1)\ s^2\ p^6\ d^{0\text{-}1}\ ns^2$ where n = 6 & 7.

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1 2
3 4 5 6 7 8

Magic Numbers

In the periodic table

Number of elements in
1st period = 2
2nd period = 8
3rd period = 8
4th period = 18
5th period = 18
6th period = 32
7th period = 32

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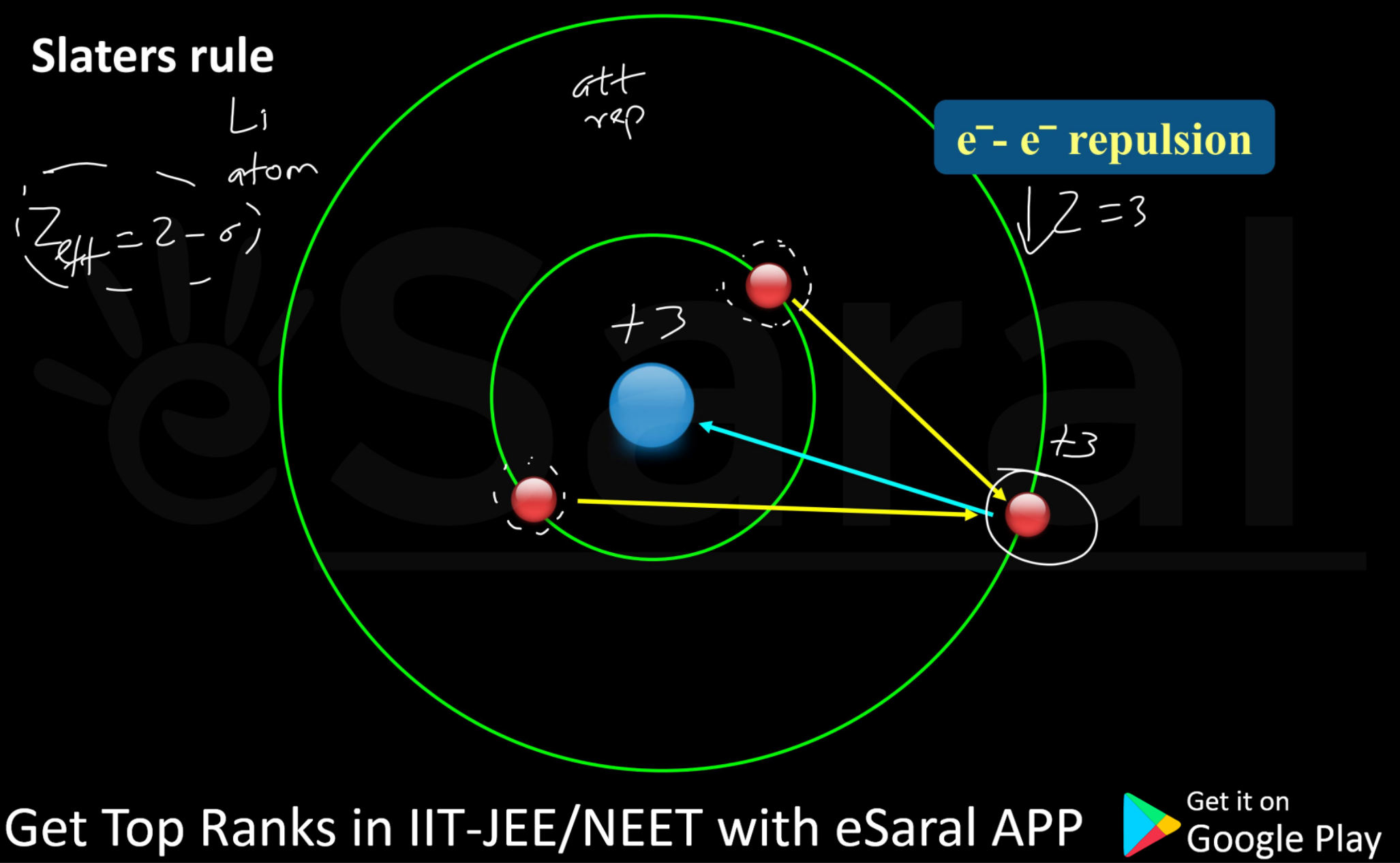
Slaters rule



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Slater's rules

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

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Calculation of σ (using Slater's rule)

$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^2$

To calculate the shielding constant (σ)

(a) Write the electronic configuration of the element in the following order and groupings :

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

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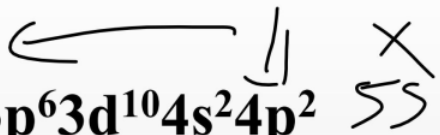

$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^2$

For s and p electrons :

(b) Electrons in any group to the right of the (ns, np) group contribute nothing to the shielding constant. (n- shell number of the electron for which σ is calculated)

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$n=4$

3×0.35

$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^2$

↑

He

$1s^2$

$2-0.3$

He $\Rightarrow Z_{eff} = 1.7$

(c) All of the other electrons in the (ns, np) group, shield the concerned electron to an extent of 0.35 each.

(Except for the 1s orbital for which value is 0.30).

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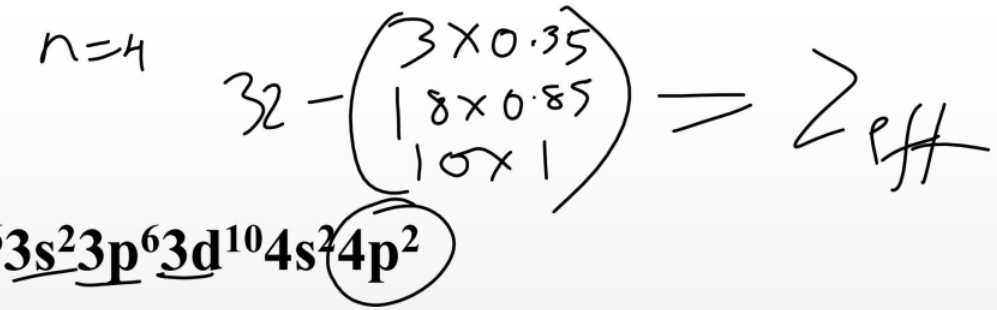

(d) All electrons in the (n – 1) shell shield to an extent of 0.85 each.

(e) All electrons (n – 2) or lower group shield completely ; that is, their contribution is 1.00 each.

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$1s^2 2s^2 2p^6 3s^2 3p^6 3d^6 4s^2$

For d and f electrons :

(a) Electrons in any group to the right of the nd or nf group contribute nothing to the shielding constant.

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$26 - (5 \times 0.35) = Z_{eff}$ for Iron

(18 × 1)

$1s^2 2s^2 2p^6 3s^2 3p^6 3d^6 4s^2$

(1)

(b) All the other electrons in the nd or nf group, shield the valence electron to an extent of 0.35 each.

(c) All electrons in groups lying to the left of the nd or nf group contribute 1.00.

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Element	Li	Be	B	C	N	O	F	Ne
Z_{eff}	1.30	1.95	2.6	3.25	3.90	4.55	5.20	5.85

From top to bottom in a group Z_{eff} remains constant for almost all.

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

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$\gamma_{shell} \propto \sigma$

For a particular n order of distance of subshells is

s < p < d < f

Hence order of shielding is

s > p > d > f

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Z_{eff} for ions

$Z_{eff} \propto +$

$Z_{eff} \propto \frac{1}{-ve}$

$Fe^{+2} > Fe$

$F^- < F$

$$Z_{eff} \propto \frac{\text{positive charge}}{\text{negative charge}}$$

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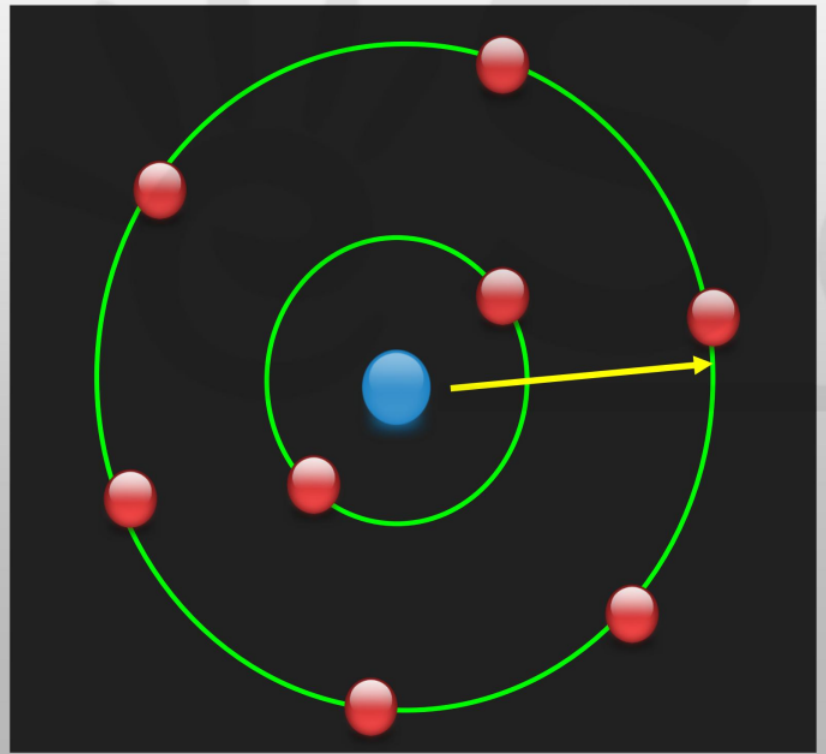
Atomic Radius

Orbitals → probability
90%.



It is very difficult to measure the atomic radius because –

(i) The isolation of single atom is very difficult.

(ii) There is no well defined boundary for the atom. (The probability of finding the e^- is 0 only at infinity).

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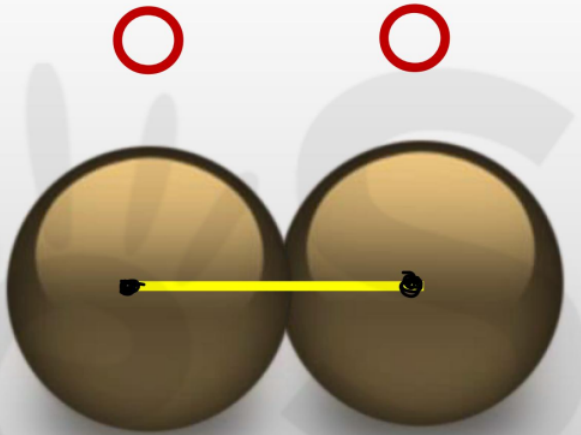

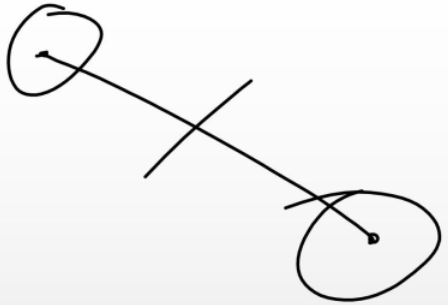

$$\text{Atomic Radius} = \frac{d_{O\text{-}O}}{2}$$

Based on the chemical bonds, atomic radius is divided into four categories –

(A) Covalent radius

(B) Ionic radius

(C) Metallic radius

(D) van der Waal's radius

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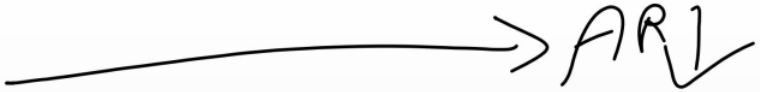


$F \uparrow \quad \frac{kq_1q_2}{r^2}$

Factors affecting atomic size are

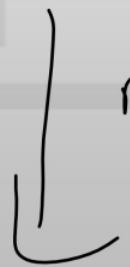
(a) Z_{eff} increases, atomic radius decreases

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

(b) Number of shell (n) increases, atomic radius increases generally

Li < Na < K < Rb < Cs

Increase in n > Z_{eff}



$n \uparrow \; AR \uparrow$

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c) Magnitude of –ve charge increases, atomic radius increases

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

d) Magnitude of +ve charge increases, atomic radius decreases

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

−ve ↑ Rep ↑ AR ↑

+ve ↑ AR ↓

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Periodic variation of atomic size

(i)Across a period (for s and p block) : It decreases from left to right in a period as effective nuclear charge (Z_{eff})increases

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

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Size ↓

Size ~

Size ↑

Nuclear > shielding

Nuclear ≈ shielding

Nuclear < shielding

Element	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn
Atomic radius (pm)	162	147	134	127	126	126	124	124	124	138

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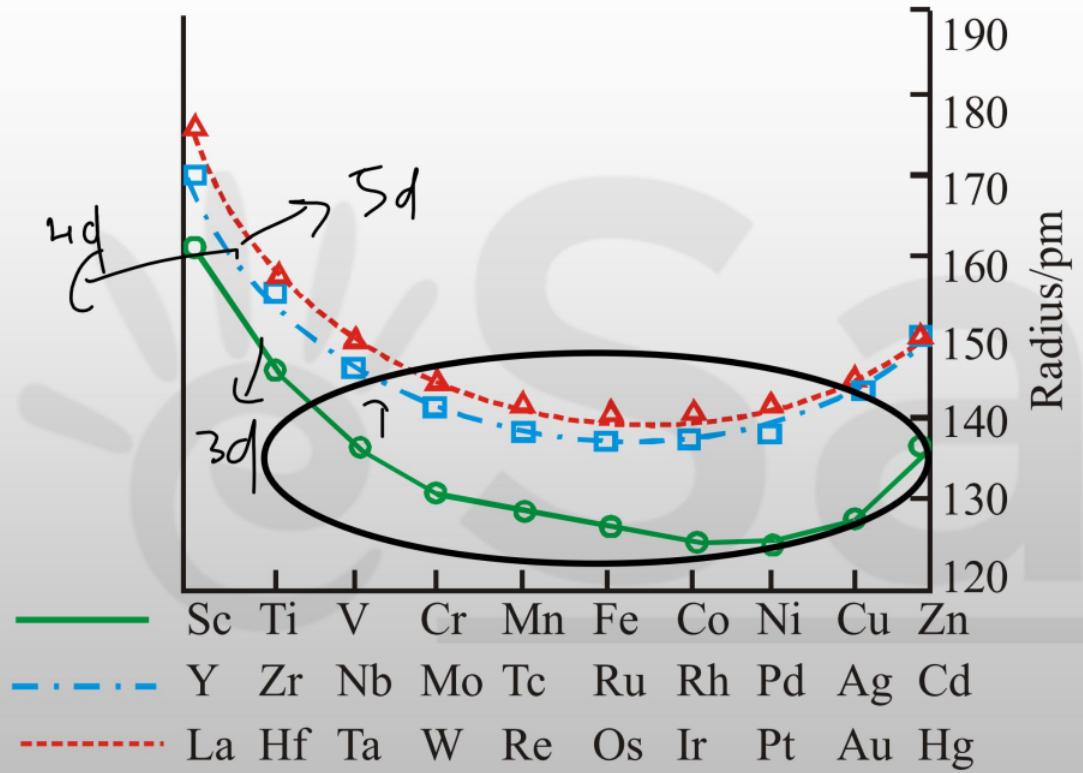

Trends in atomic radii of transition elements

ARN

4d ~ 5d

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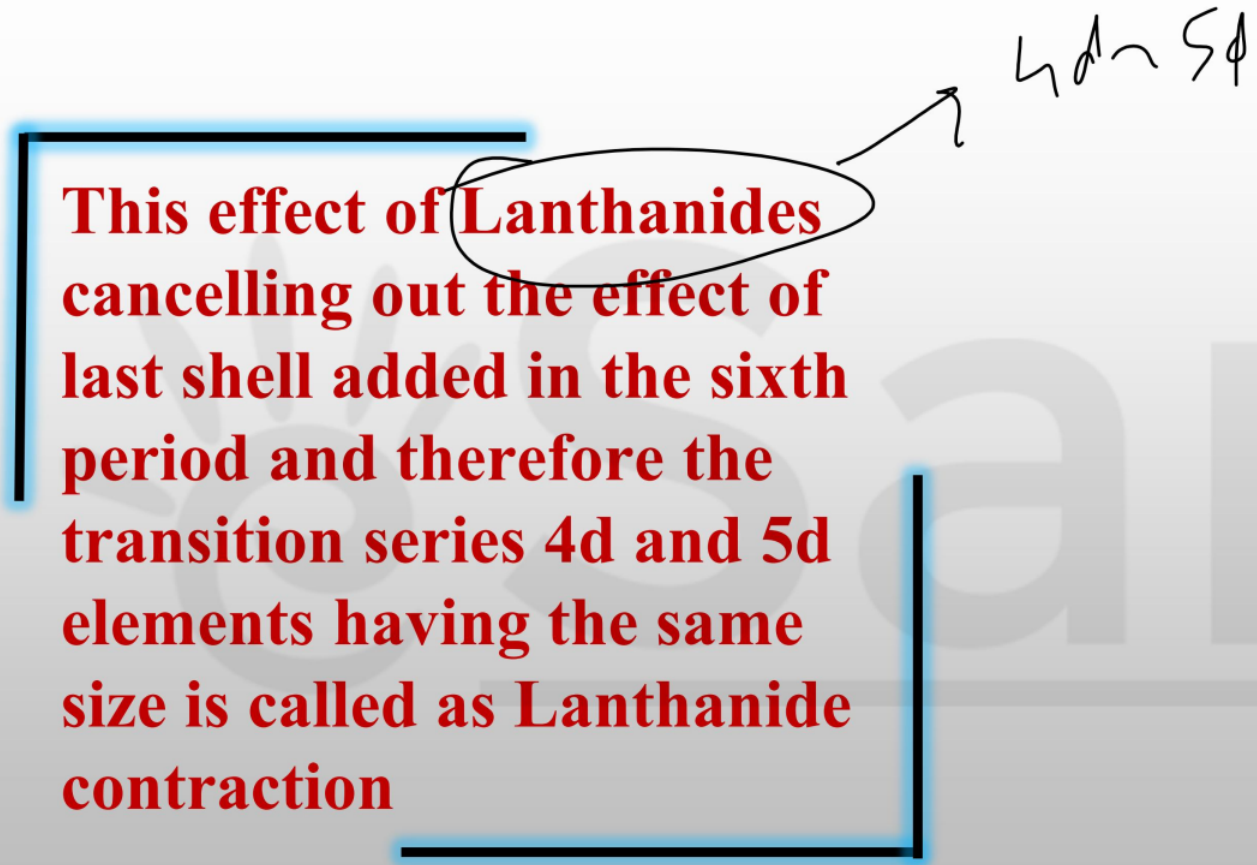

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IONISATION ENERGY

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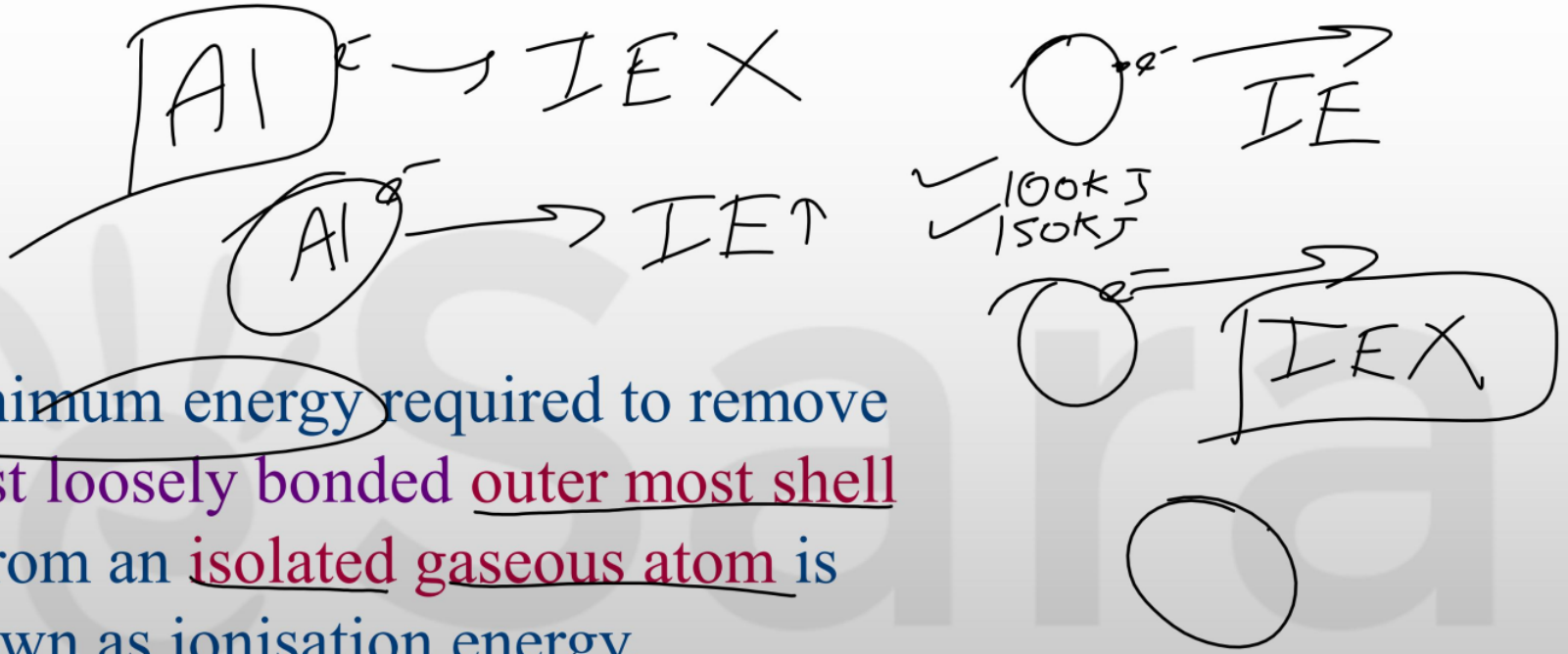
Ionisation Potential / Ionisation Energy / Ionisation Enthalpy (IP / IE)



Minimum energy required to remove most loosely bonded outer most shell e^- from an isolated gaseous atom is known as ionisation energy.

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Successive Ionisation Energy

$O \rightarrow O^-$

$Z_{eff}\uparrow$

$$M_{(g)} + E_1 \xrightarrow{E_2} M^+_{(g)} + e^-$$

$$M^{+1}_{(g)} + E_2 \xrightarrow{E_1} M^{+2}_{(g)} + e^- \quad (Z_{eff}\uparrow)$$

$$M^{+2}_{(g)} + E_3 \xrightarrow{E_3} M^{+3}_{(g)} + e^-$$

E_1 = Ist Ionisation energy

E_2 = 2nd Ionisation energy

E_3 = 3rd Ionisation energy

$$E_1 < E_2 < E_3 \ldots\ldots\ldots$$

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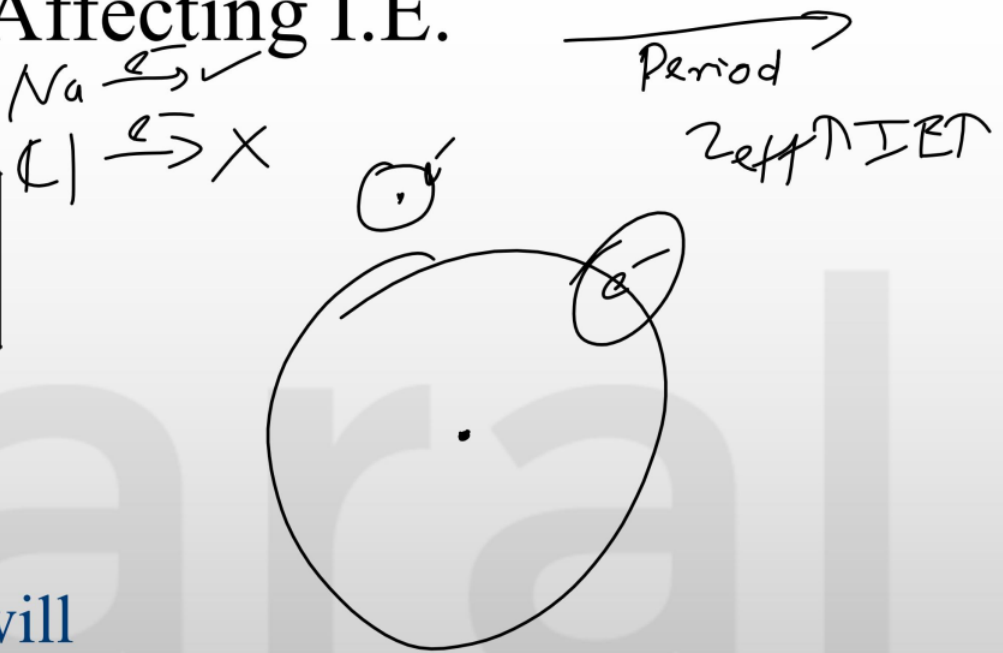
Factors Affecting I.E.

$$\text{Ionisation Energy} \propto \frac{1}{\text{atomic size}}$$



$$\text{Ionisation Energy} \propto Z_{eff}$$

Ion with higher oxidation state will have more ionisation energy.

Ex. $Fe^{+3} > Fe^{+2} > Fe$

Penetration power of sub shells

$$s > p > d > f$$

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(vi) Stability of half filled and fully filled orbitals :

Half filled p^3,d^5, f^7 or fully filled s^2, p^6, d^{10}, f^{14} are more stable than others so it requires more energy to remove an electron.

IE_1 N > O

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

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Exception :

Learn

Ionisation Energy Ga > Al (due to poor shielding of 3d)

Ionisation Energy of 5d > 4d (due to lanthanide contraction)

Ex. Hf > Zr

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(B) Reactivity of metals

$Cs > Rb > K > Na$

$$\text{Reactivity of metals} \propto \frac{1}{\text{IE}}$$

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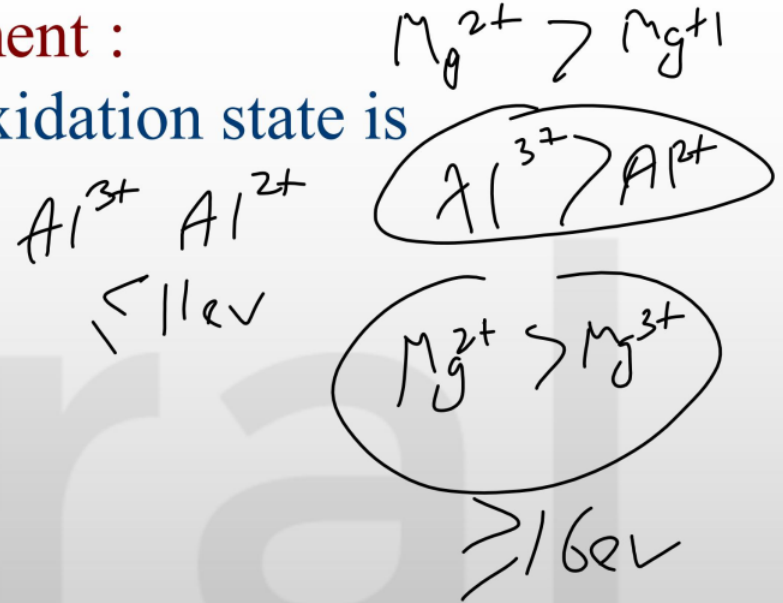
(C) Stability of oxidation states of an element :

(a) successive IE $\geq$ 16eV, then its lower oxidation state is stable.

(b) $\leq$ 11 eV then higher oxidation state



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Q) The IP_1, IP_2, IP_3, IP_4 and IP_5 of an element are 7.1, 14.3, 34.5, 46.8, 162.2 eV respectively. The element is likely to be:-

(1) Na (2) Si (3) F (4) Ca

huge jump → noble gas

Ans 2

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Electron Gain Enthalpy (ΔH_{eg})

(1) The amount of energy change when an electron is added to the valence shell of an isolated gaseous atom is known as EGE.

$$X(g) + e^- \longrightarrow X^-(g) \qquad \Delta H_{eg}$$

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Is EGE positive or negative?

3 exception

$EGE_I \Rightarrow -ve$

Generally first electron addition in an isolated gaseous atom is an exothermic process
(except stable electronic configuration)

$EGE_{II} \Rightarrow +ve$ all

$$X(g) + e^- \longrightarrow X^-(g) + EGE$$

Second (and so on) electron addition in an isolated gaseous atom is always an endothermic process due to inter electronic repulsion.

$$X^-(g) + e^- \longrightarrow X^{-2}(g)$$

ΔH_{egII} = positive

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$$\text{Atomic Size} \propto \frac{1}{EA}$$

EGE = − EA

− 10

− 150

→ |EGE|

EGE of N, Be, Ne are positive

In period – EGE becomes more negative generally as we go from left to right in a period.

Li *Be* B C *N*

O F *Ne*

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F < Cl

Yahi se aayega

Imp

(b) In Group

In a group, the electron affinity decreases on moving from top to bottom, that is, less and less amount of energy is released.

Electron affinity of 3rd period element is greater than electron affinity of 2nd period elements of the respective group.

F	Cl
[He] $2s^2 2p^5$	[Ne] $3s^2 3p^5$

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Electronegativity (χ) :-

no exception

Period → EN↑

↓ EN ↓

The tendency of a covalently bonded atom to attract shared pair of electrons towards itself is called electronegativity.

I⊖Cl →

F — H ←

1. It is not an energetic term.
2. It regularly increases in a period because it does not depend upon stable electronic configuration
3. It has no unit
4. It is a relative term

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F = 4

O = 3.5

N = 3

Cl = 3

P, M, Cm

1) Mulliken – Jaffe
2) Alfred – Rochow
3) Pauling Scale

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Factors Affecting electronegativity

$EN \propto Z_{eff}$ $\quad$ +ve ↑ EN

(a) Charge on atom:

The cation will be more electronegative than parent atom.

As the +ve charge on the cation increases electronegativity increases.

E.N. of $Fe^{+3} > Fe^{+2} > Fe$

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The anion will be less electronegative than atom. As the negative charge on the anion increases electronegativity decreases.

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

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$$\text{Electronegativity} \propto \frac{1}{\text{Atomic size}}$$

H — H

H — X ↑

Ex. F > Cl > Br > I

$$\text{Electronegativity} \propto Z_{eff} \propto \frac{\text{positive charge}}{\text{negative charge}}$$

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Variation in Electronegativity

(a) Electronegativity decreases down the group.

(b) In period on moving from left to right electronegativity increases.

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Applications of E.N.

EN NMT

Metallic and non metallic properties of elements

(i) The metallic character decreases as the electronegativity of the element increases.

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(ii) On moving from left to right in a period, the electronegativity of the elements increases. So the metallic character decreases.

(iii) On moving down a group, the electronegativity of the elements decreases, so the metallic character increases.

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Schomaker and Stevenson law

A → 10

B → 8

$d_{A-B} = 10 + 8 - 0.09(1)$

$= 17.91$

As per Schomaker and Stevenson–
The reduction in bond length depends on the difference in electronegativities of atoms in following manner –

$$d_{A-B} = r_A + r_B - 0.09(\chi_A - \chi_B)$$

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(iv) Nature of oxides

$Na_2O \rightarrow$ basic SO_3 acidic

$N_2O_5 \rightarrow$ acidic

Acidic Nature $\propto$ Oxidation state

1. Along a period acidic nature increases.
2. Down the group basic nature increases

$\overset{+6}{SO_3} > \overset{+4}{SO_2}$

$\overset{+5}{N_2O_5} > \overset{+3}{N_2O_3}$

OSP ANT

NaOH

CsOH

Li	Be	B	C	N	O	F
Na	Mg	Al	Si	P	S	Cl

Basic Amphoteric Acidic

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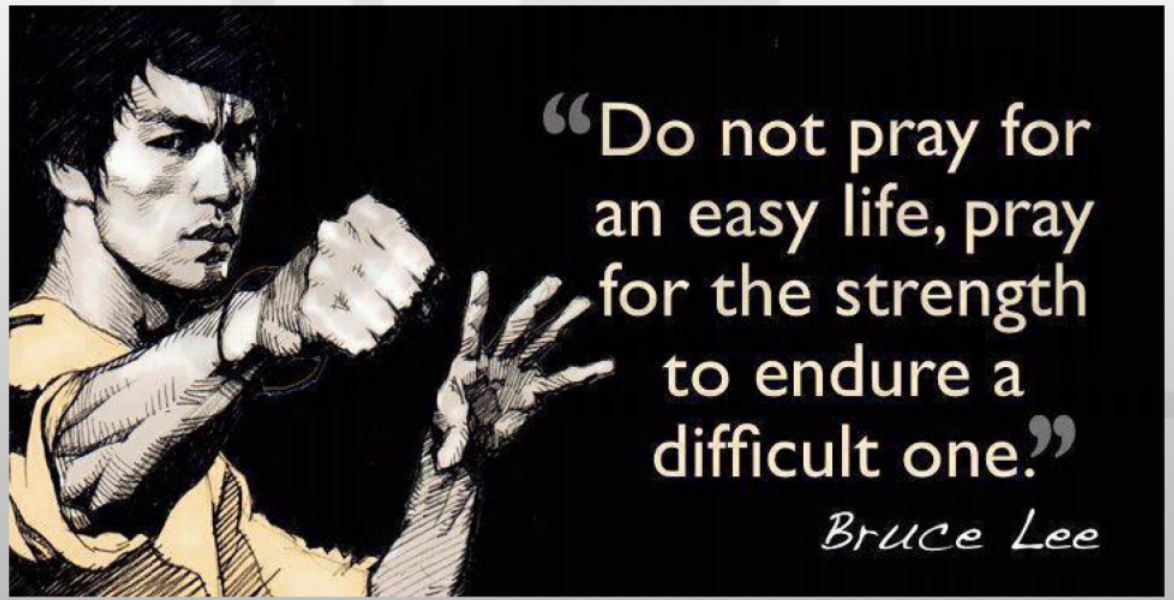
musibat

hurdles

problems



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5 B
13 Al
31 Ga
49 In
81 Tl

The Boron Family

Bengan Aloo Gajar in Thela

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i. Boron is a typical non-metal.

ii. Aluminum is a metal but shows many chemical similarities to Boron and Gallium.

iii. Indium and Thallium are almost exclusively metallic in character.

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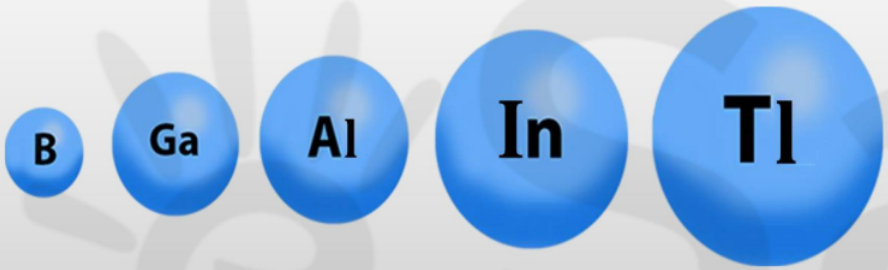
Atomic and Ionic radii order

B < Ga < Al < In < Tl

13 Al → [Ne] $3s^2\ 3p^1$

31 Ga → [Ar] $3d^{10}\ 4s^2\ 4p^1$

Extra



Atomic radius of Ga is less than that of Al due to poor shielding of 3d electrons in Ga.

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Boiling point order

BP ∝ MM

B.P. B > Al > Ga > In > Tl

Group 13 ↓ BPt

Melting point order

M.P. B > Al > Tl > In > Ga

Ionization Enthalpy

B > Tl > Ga > Al > In

Electronegativity

B > Tl > In > Ga > Al

Density of the elements increases down the group from boron to thallium.

B < Al < Ga < In < Tl

↓ of T

Electropositive Character

$$\frac{\text{B}}{\text{Non metal}} < \frac{\text{Al} < \text{Ga} < \text{In} < \text{Tl}}{\text{Metals}}$$

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Chemical Properties

B^{+3} ✗ Boron rarely forms ionic compound

(i) Due to small size of boron, the sum of its first three ionization enthalpies is very high. This prevents it to form +3 ions and forces it to form only covalent compounds.

Tl +1 > Tl +3 → Inert pair effect

As a result of this, only p-orbital electron may be involved in bonding. In fact in Ga, In and Tl, both +1 and +3 oxidation states are observed.

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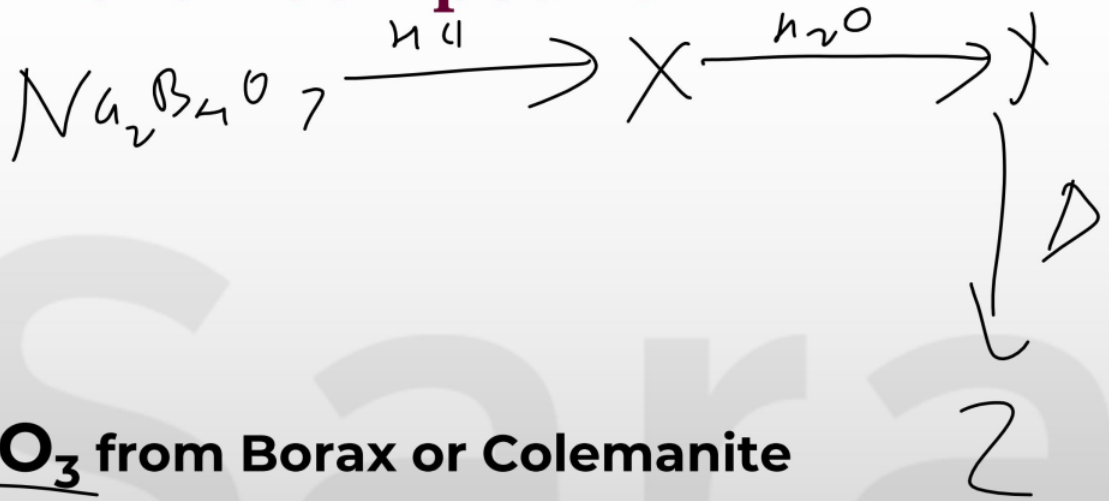
Preparation of Boron compound



(i) Preparation of B_2O_3 from Borax or Colemanite

$$Na_2B_4O_7 + HCl/H_2SO_4 \longrightarrow NaX + H_2B_4O_7$$

$$H_2B_4O_7 + 5H_2O \longrightarrow 4H_3BO_3 \xrightarrow{\Delta} B_2O_3 + H_2O$$

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Borax

Preparation of Borax

$2CaO.\ 3B_2O_3 + 2Na_2CO_3 \longrightarrow 2CaCO_3\downarrow + Na_2B_4O_7 + 2NaBO_2$

Colemanite

Filtered
- $CaCO_3$(as residue)

$Na_2B_4O_7 + NaBO_2$

in solution

Concentrated and allowed to crystallise out and filtered

$NaBO_2$ + $Na_2B_4O_7 \cdot 10H_2O$

in filtrate as residue

CO_2 Passed and crystallise out again

$[4NaBO_2 + CO_2 \longrightarrow Na_2B_4O_7 + Na_2CO_3]$

$Na_2B_4O_7 . 10H_2O\downarrow$

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$$Na_2B_4O_7.10H_2O \xrightarrow{\Delta} Na_2B_4O_7 \xrightarrow{\Delta} 2NaBO_2 + B_2O_3$$

WW Imp

Sodium metaborate

Boric anhydride

[Heating effect]

$Na_2B_4O_7 \xrightarrow{\Delta} X + B_2O_3$

$\xrightarrow{\Delta} B_2O_3 + X$

(ii) Borax dissolves in water to give an alkaline solution.

$$Na_2B_4O_7 + 7H_2O \longrightarrow 2NaOH + 4H_3BO_3$$

Orthoboric acid

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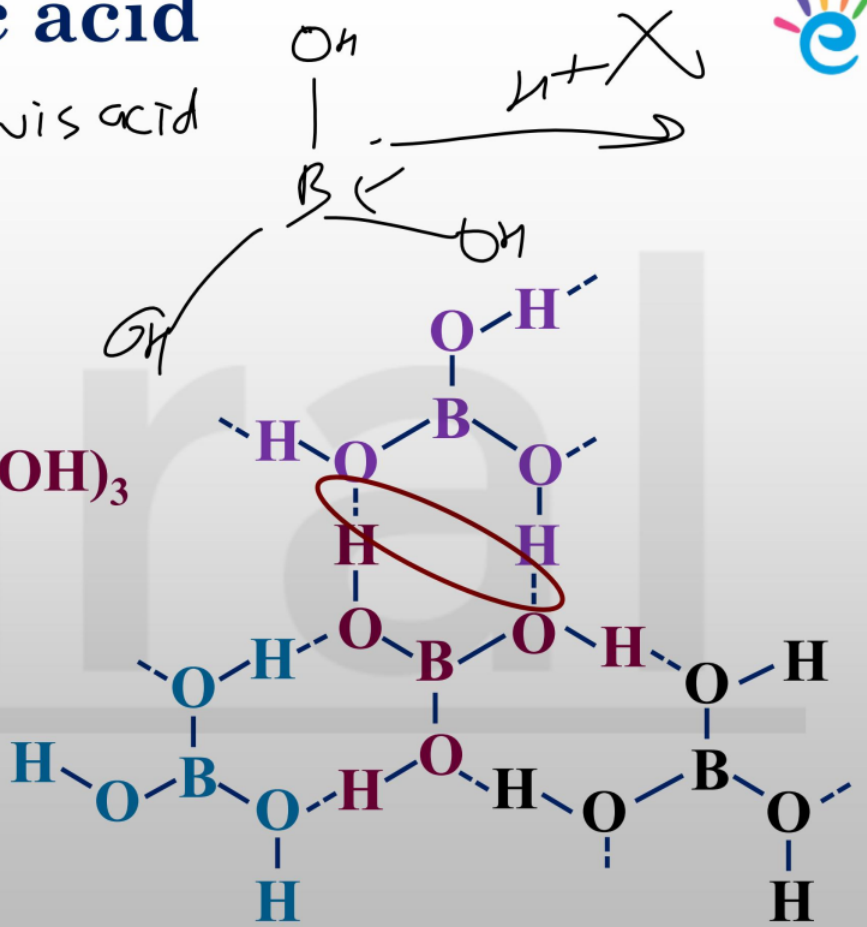
Orthoboric acid

Preparation

(i) It can be prepared by acidifying an aqueous solution of borax.

$$Na_2B_4O_7 + 2HCl + 5H_2O \longrightarrow 2NaCl + 4B(OH)_3$$

Basicity 1



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Heating of Boric acid

$$H_3BO_3 \xrightarrow{100^\circ C} HBO_2 \xrightarrow{140^\circ C} H_2B_4O_7 \xrightarrow[\text{red hot}]{\Delta} B_2O_3$$

Metaboric acid Tetra boric acid Glassy mass

OBA

$\xrightarrow{\Delta}$ B_2O_3

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Diborane, B_2H_6

Preparation

Structure

3c-2e⁻

The simplest boron hydride known as diborane.

$$3LiAlH_4 + 4BF_3 \longrightarrow 3LiF + 3AlF_3 + 2B_2H_6$$

or $LiBH_4$ **or $3(BF_3)$**

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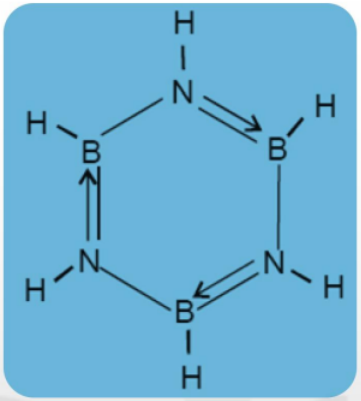

Reaction of ammonia with diborane gives initially $B_2H_6.2NH_3$ which is formulated as $[BH_2(NH_3)_2]^+ [BH_4]^-$;

further heating gives borazine, $B_3N_3H_6$ known as "inorganic benzene" in view of its ring structure with alternate BH and NH groups.

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Al_2O_3 Preparation

Aluminium

(i) $2Al(OH)_3 \xrightarrow{300°C} Al_2O_3 + 3H_2O$

↳ metallurgy

(ii) $Al_2(SO_4)_3 \xrightarrow{\Delta} Al_2O_3 + 3SO_3$

Alum

(iii) $(NH_4)_2SO_4 \cdot Al_2(SO_4)_3 \cdot 24H_2O \xrightarrow{\Delta}$

$Al_2O_3 + 2NH_3 + 4SO_3 + 25H_2O$

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$AlCl_3$ Preparation

$$2Al + 6HCl\,(vap.) \longrightarrow 2AlCl_3 + 3H_2$$

(over heated) dry

$$Al_2O_3 + 3C + 3Cl_2 \xrightarrow[1000^\circ C]{\Delta} 2AlCl_3\,(vap.) + 3CO$$

Cooled ↓

Solid anh. $AlCl_3$

(i) Its anhydrous form is deliquescent and fumes in air.

(ii) It sublimes at 180°C.

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Alumns

no of water of hydration in Potash alum

$M_2SO_4 . M'_2(SO_4)_3 \cdot 24\ H_2O$

where

$M = Na^+, K^+, Rb^+, Cs^+, As^+, Tl^+,$

$M' = Al^{+3}, Cr^{+3}, Fe^{+3}, Mn^{+3}, Co^{+3}$

$K_2SO_4 \cdot Al_2(SO_4)_3 \cdot 24H_2O$ Potash alum

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Uses

(i) Acts as coagulant

(ii) Purification of water

(iii) Tanning of leather

(iv) Mordant in dyeing

(v) Antiseptic

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The Carbon Family

Chemistry Sir Gives Sanki Problems

5 C	→	$1s^2\ 2s^2\ 2p^2$
13 Si	→	$[Ne]\ 3s^2\ 3p^2$
31 Ge	→	$[Ar]\ 3d^{10}\ 4s^2\ 4p^2$
49 Sn	→	$[Kr]\ 4d^{10}\ 5s^2\ 5p^2$
81 Pb	→	$[Xe]\ 4f^{14}\ 5d^{10}\ 6s^2\ 6p^2$

$ns^2\ np^2$ ↑↓ ↑ ↑ ↑

4

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Covalent Radius

Covalent radii C < Si < Ge < Sn < Pb

n↑ size↑

Ionization Enthalpy

Decreases down the group

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Melting and Boiling Points

M.P. C > Si > Ge > Pb > Sn

B.P. C > Si > Ge > Sn > Pb

Electronegativity

The electronegativity values for elements from Si to Pb are almost the same.

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Reactivity Towards Oxygen

All members when heated in oxygen form oxides.

There are mainly two types of oxides, i.e., monoxide and dioxide of formula MO and MO_2 respectively.

SiO only exists at high temperature.

The dioxides — CO_2, SiO_2 and GeO_2 are acidic, whereas SnO_2 and PbO_2 are amphoteric in nature.

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Reactivity Towards Halogen

(i) These elements can form halides of formula MX_2 and MX_4 (where X = F, Cl, Br, I).

(ii) Except carbon, all other members react directly with halogen.

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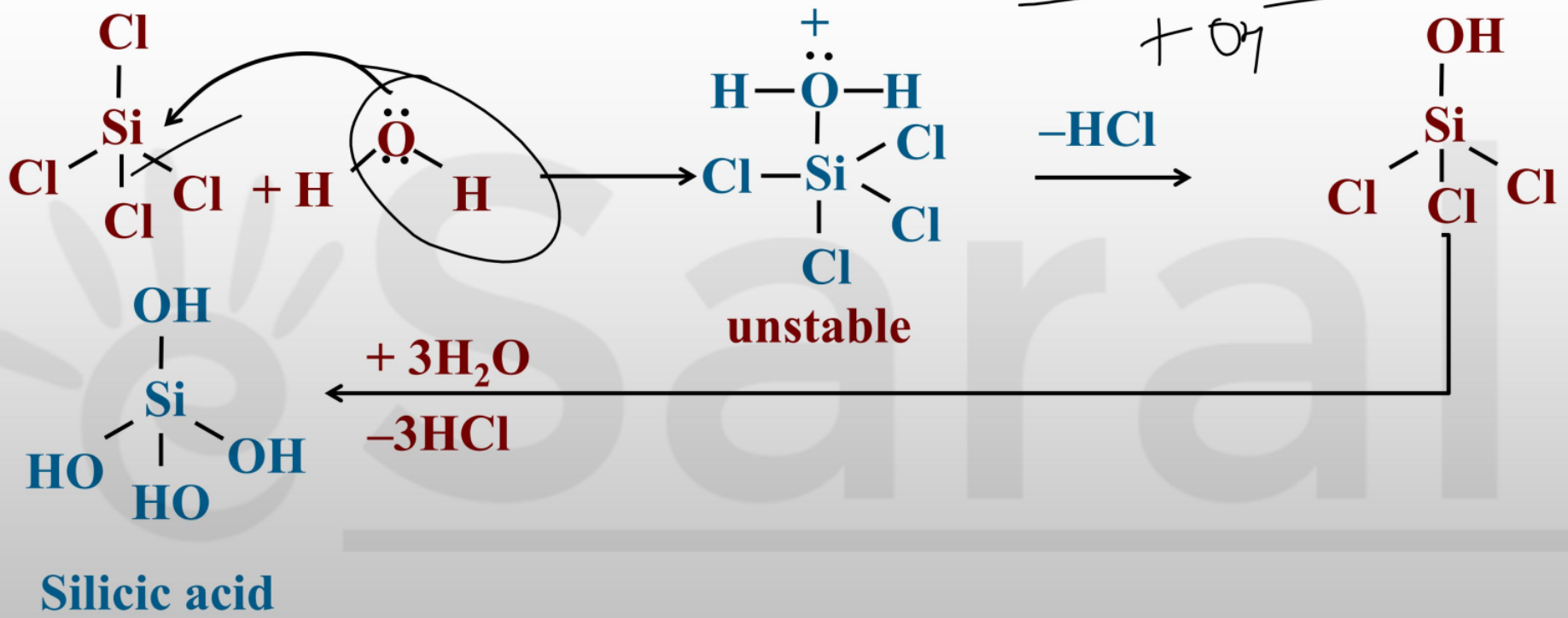
Hydrolysis of $SiCl_4$ (Mechanism)

VVImp — Cl

+ OH



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Allotropes of Carbon

Graphite — Lubricant

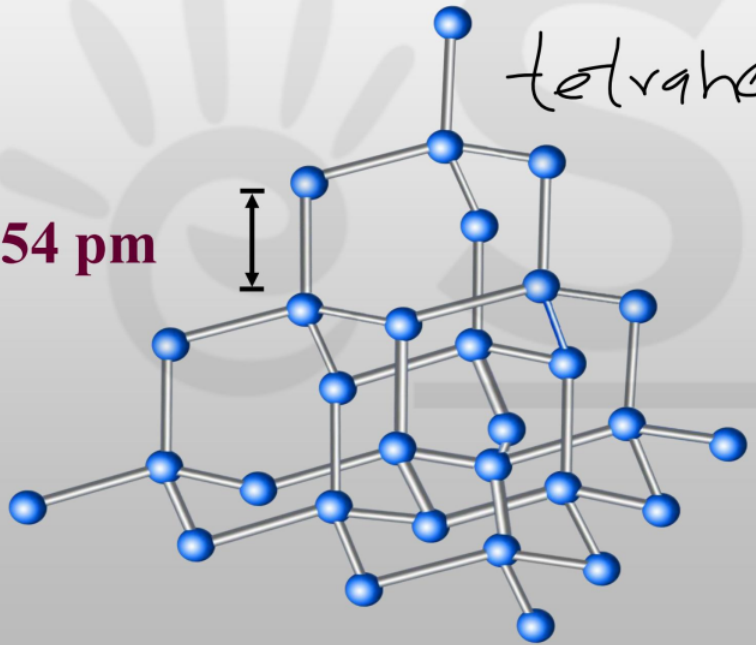
Imp

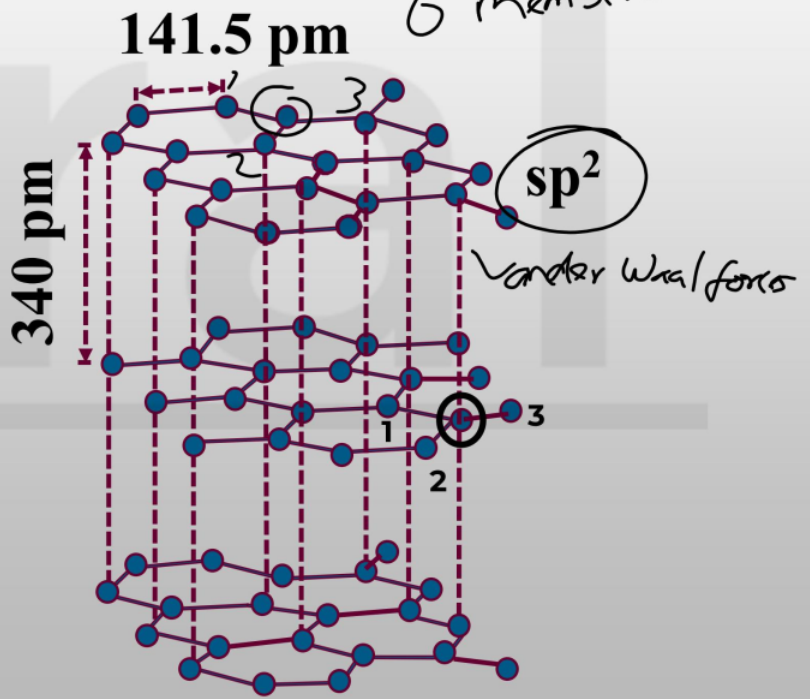
Stable

thermodynamically

6 membered

tetrahedral





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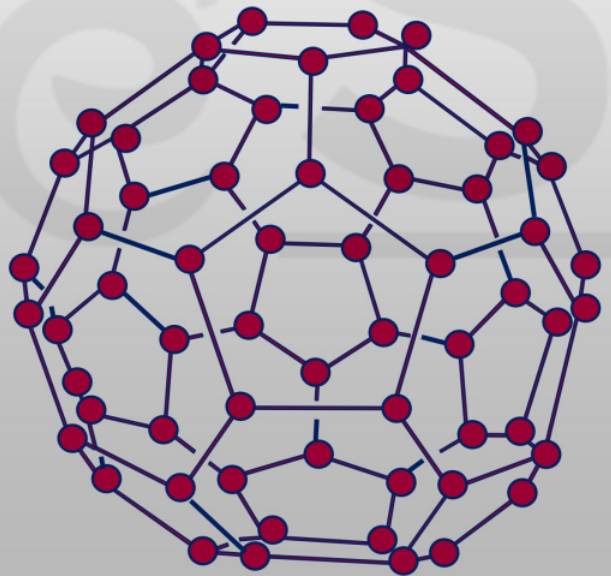
Fullerenes

(i) Fullerenes are made by the heating of graphite in an electric arc in the presence of inert gases such as helium or argon.



C_{60}

sp^2

5 & 6 membered rings

60 to 350 Carbon rings

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Carbide

C_2^{2-}

Types of Carbide

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(i) Ionic and salt like

Classification on basis of no. of carbon atoms present in hydrocarbon found on their hydrolysis.

(a) C_1 unit

(b) C_2 unit

(c) C_3 unit

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C^{4-}

C_1 unit Be_2C , Al_4C_3

$Be_2C + 4H_2O \longrightarrow 2Be(OH)_2 + CH_4$

C_2 unit CaC_2, BaC_2

$CaC_2 + 2H_2O \longrightarrow Ca(OH)_2 + CH \equiv CH$

C_3 unit Mg_2C_3

$Mg_2C_3 + 4H_2O \longrightarrow 2Mg(OH)_2 + CH_3 - C \equiv CH$

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(ii) Covalent carbide

CH_4, CO_2, CS_2 can be considered as covalent carbide and SiC & B_4C also consider as covalent carbide.

(iii) Interstitial carbide

Transition element or inner transitional elements forms this kind of carbide.

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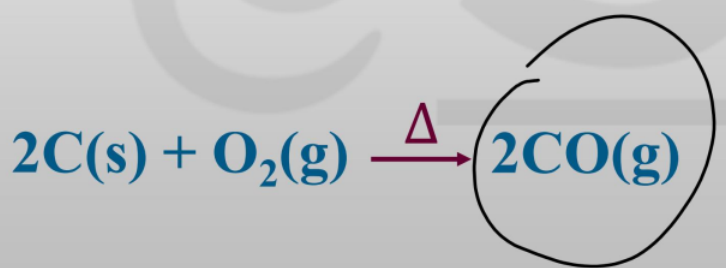
Carbon Monoxide

Poisonous

CO_2 ✗ CO

Prepration

(i) Direct oxidation of C in limited supply of oxygen or air yields carbon monoxide.



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Poisonous Nature of CO

The highly poisonous nature of CO arises because of its ability to form a complex with hemoglobin, which is about 300 times more stable than the oxygen-hemoglobin complex

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Carbon Dioxide

Prepration

(i) It is prepared by complete combustion of carbon and carbon containing fuels in excess of air.

$$C(s) + O_2(g) \xrightarrow{\Delta} CO_2(g)$$

$$CH_4(g) + 2O_2(g) \xrightarrow{\Delta} CO_2(g) + 2H_2O(g)$$

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(iii) With water, it forms carbonic acid, H_2CO_3 which is a weak dibasic acid and dissociates in two steps

$$H_2CO_3(aq) + H_2O(l) \rightleftharpoons HCO_3^-(aq) + H_3O^+(aq)$$

$$HCO_3^-(aq) + H_2O(l) \rightleftharpoons CO_3^{2-}(aq) + H_3O^+(aq)$$

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Use of CO_2

Photosynthesis.

$$6CO_2 + 12H_2O \xrightarrow[\text{Chlorophyll}]{hv} C_6H_{12}O_6 + 6O_2 + 6H_2O$$

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Harmful Effect of CO_2

Unlike CO, it is not poisonous.

But the increase in combustion of fossil fuels and decomposition of limestone for cement manufacture in recent years seem to increase the CO_2 content of the atmosphere.

This may lead to increase in green house effect and thus, raise the temperature of the atmosphere which might have serious consequences.

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Silicon (Si)

Silica is found in the free state in sand, flint and quartz and in the combined state as silicates like

(i) Feldspar $K_2O.\ Al_2O_3.\ 6SiO_2$

(ii) Kaolinite $Al_2O_3.\ 2SiO_2.\ 2H_2O$

(iii) Asbestos $CaO.\ 3MgO.\ 4SiO_2$

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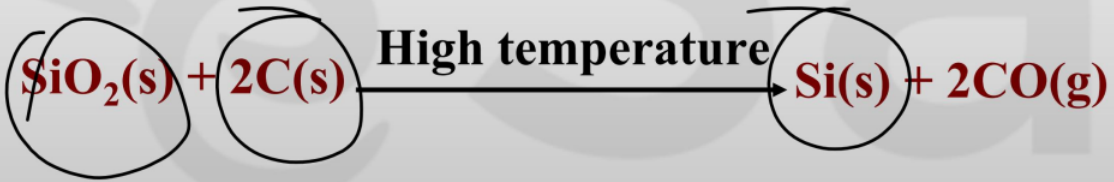
Preparation

(i) From silica (sand)

Elemental silicon is obtained by the reduction of silica (SiO_2) with high purity coke in an electric furnace.



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Chemical Properties

Si → Catenation

Silicon is particularly unreactive at room temperature towards most of the elements except fluorine.

$$Si(s) + 2F_2(g) \xrightarrow{\text{Room Temperature}} SiF_4(l)$$

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Compounds of Silicon

Silane

$Si_nH_{2n+2} \rightarrow SiH_4$ & Si_2H_6

Only these two are found

Silicones

Linear chain compound

(ii) Linear Silicones

$$R_2SiCl_2 + H_2O \xrightarrow[-2HCl]{} R_2Si(OH)_2 \xrightarrow[-H_2O]{\Delta} -O-\underset{\underset{R}{|}}{\overset{\overset{R}{|}}{Si}}-O-\underset{\underset{R}{|}}{\overset{\overset{R}{|}}{Si}}-O-\underset{\underset{R}{|}}{\overset{\overset{R}{|}}{Si}}-O-$$

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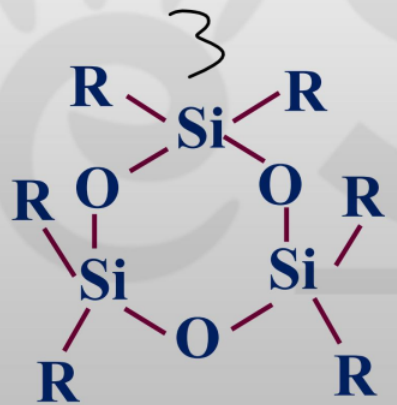
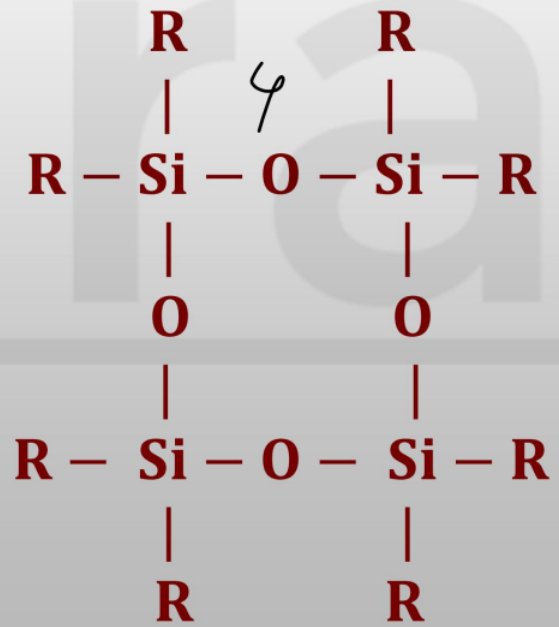
(iii) Cyclic Silicones

Silicones may have the cyclic structure also having 3, 4, 5 and 6 silicon atoms within the ring.





cyclic silicones are non planar

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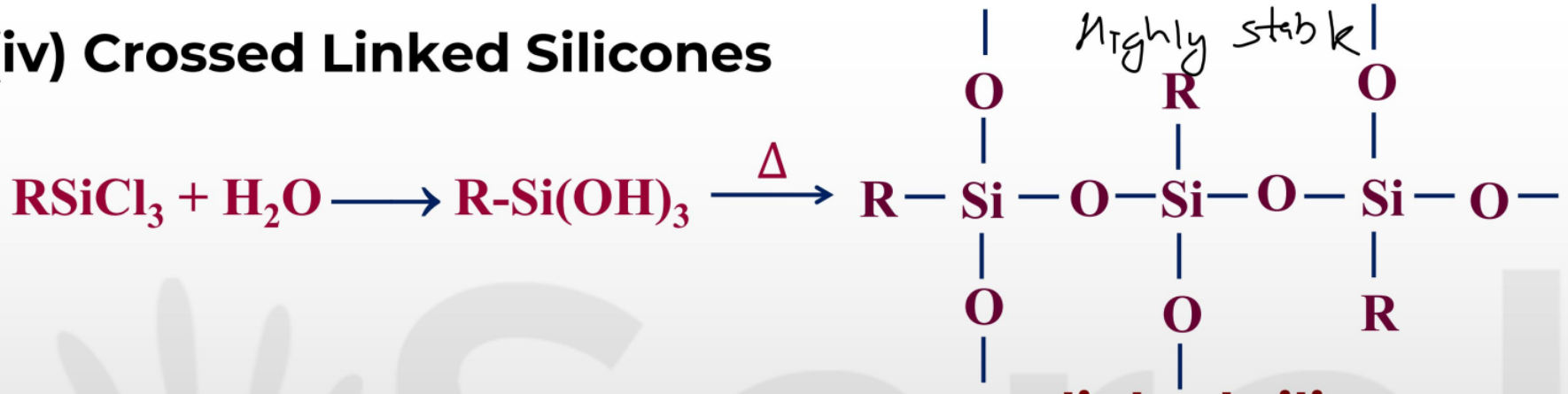
(iv) Crossed Linked Silicones



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Silicates

Silicates are metal derivatives of silicic acid, H_4SiO_4 or $Si(OH)_4$.

Silicates are formed by heating metal oxide or metal carbonates with sand, e.g.,

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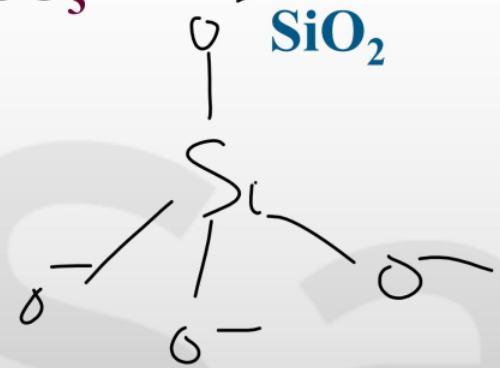
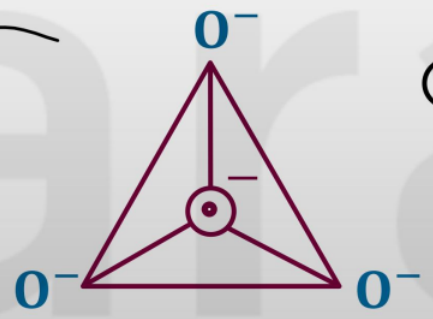
$$Na_2CO_3 \xrightarrow{\text{Fused with sand}} Na_4SiO_4,\ Na_2(SiO_3)_n,\ \text{etc.}$$

SiO_2





O → Oxygen

• → Silicon

O⁻ Si O⁻ O⁻

Plane projection of silicate ion

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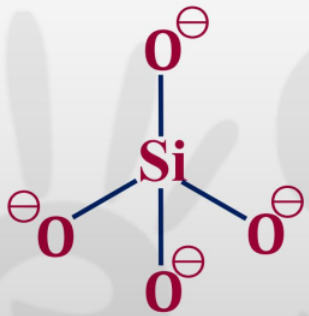
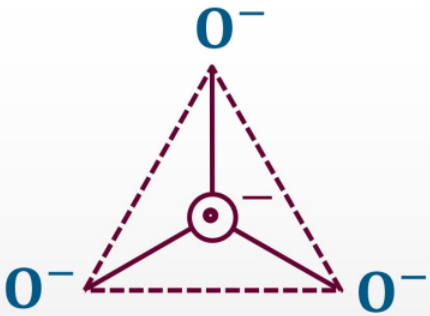
(1) Ortho silicates (Ex. Zircon, Willemite)

(2) Minerals in which SiO_4^{-4} anion is present in discrete form (not in polymeric form) are called Orthosilicate.





• ⟶ Silicon

○ ⟶ Oxygen

Si = sp^3 hybridisation

Tetrahedral in shape

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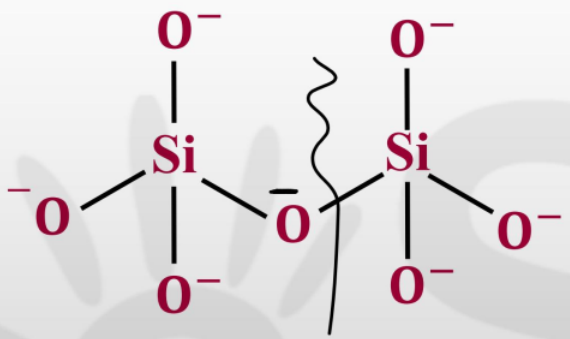
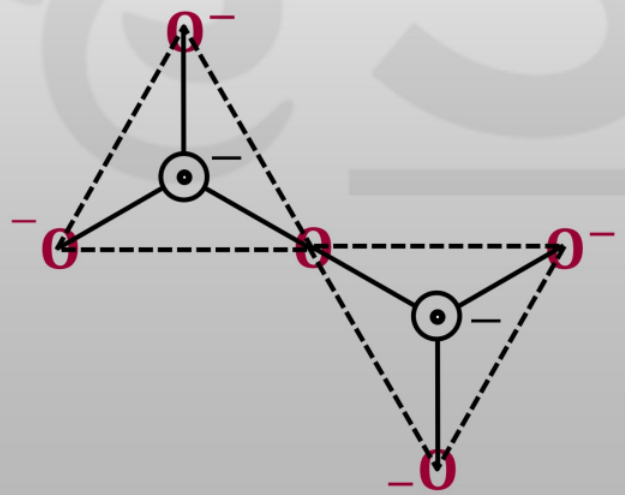
(2) Pyro Silicates (Ex. Hemimorphite)

O^{2-}

$Si_2O_7^{6-}$





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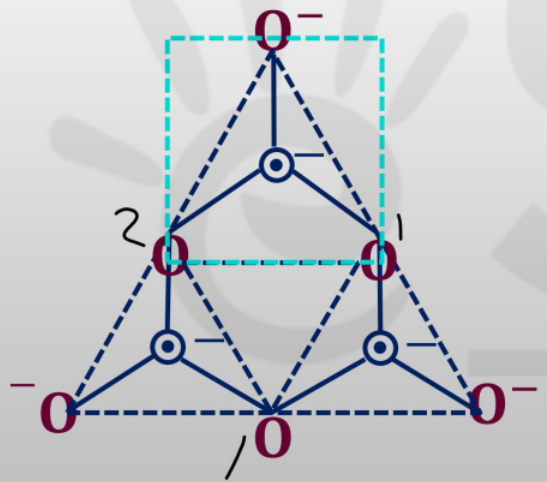
(3) Cyclic Silicate (Wollastonile, Benitotite, Beryl, Emrald)

Shared O → 0

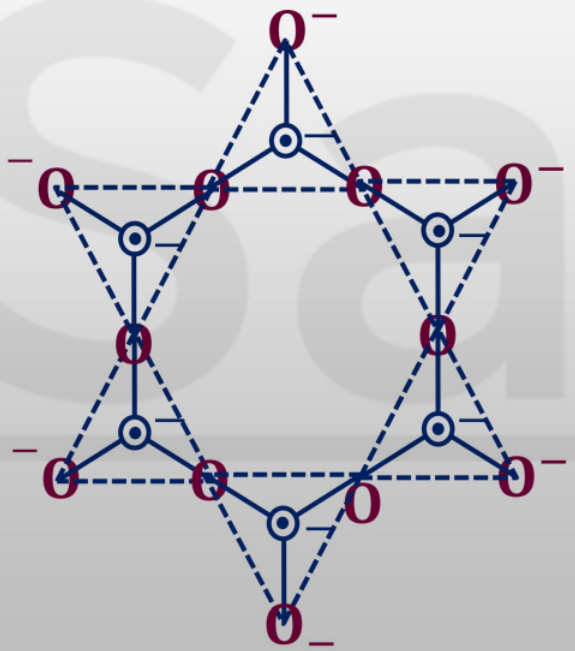
non-shared → -1

$Si_nO_{3n}^{2n-}$

$Si_3O_9^{-6}$



$Si_6O_{18}^{-12}$



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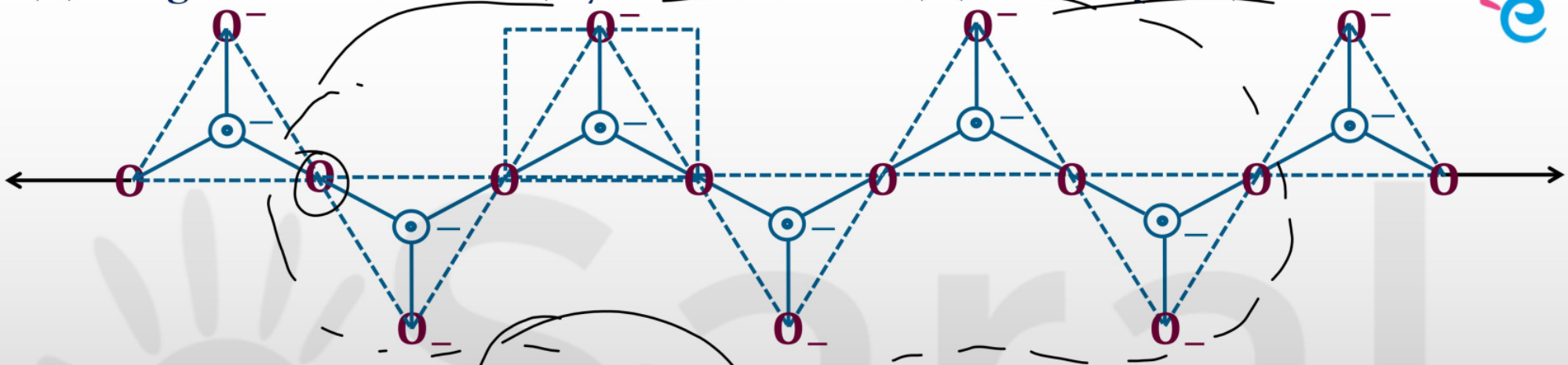
(4) Single chain silicate (Pyroxene Silicate) (Ex. Diopsite)



Formula = $(SiO_3)_n^{-2n}$ charge on one unit = –2

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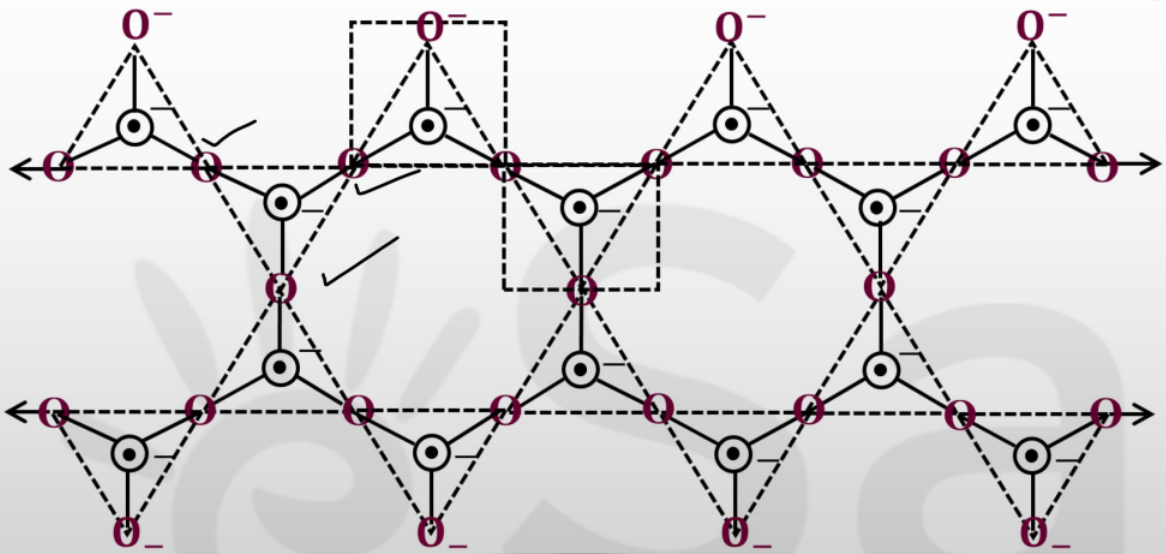
(5) Double chain Silicate (Amphiboles) (Ex. Asbestos)

VVImp



Formula = $(Si_4O_{11})_n^{-6n}$

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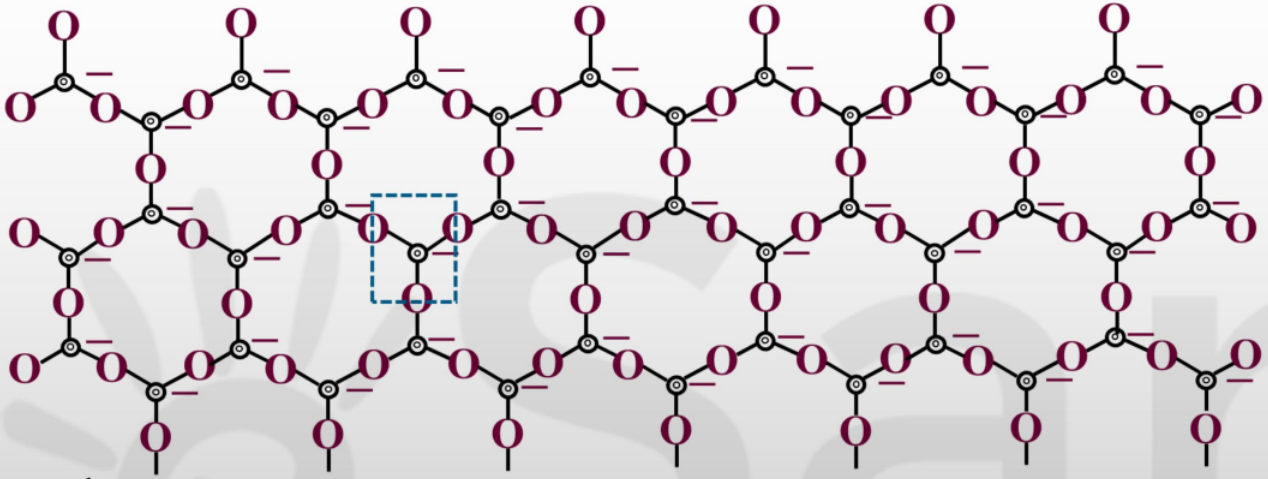
(6) Sheet Silicate (Ex. Clay talc, Micas)

General Form

Examples



Gen. Formula $(SiO_{2.5})_n^{-n}$ or $(Si_2O_5)_n^{-2n}$

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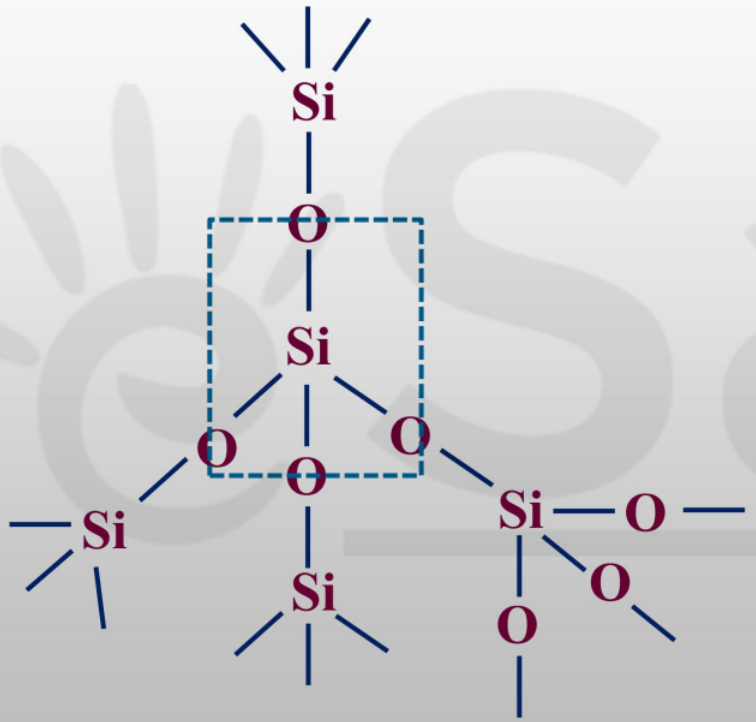
(7) 3-D Silicate (Ex. fledspars, zeolites, ultramarine, quartz)

non-ionic



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