ARJUNA NEET BATCH

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Classification of Elements & Periodicity in Properties

LECTURE-05

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Quick Revision:

Atomic size

$r_{vander\,waal} > r_{metallic} > r_{covalent}$

Trend

PERIOD

no. of shells : same

Z_{eff} ↑

size ↓ (till Gp17 & then ↑ses to Gp18)

GROUP

no. of shells ↑

size ↑

Exceptions:

① Transition elements
   ↳ 3d series

② Lanthanoid contraction

Size: $3d < 4d \approx 5d$ (except Gp IIIB)

eg. Zr & Hf have similar radii.

③ Transition contraction:

$B < Al > Ga$ or $B < Ga < Al$

Ga ↳ 3d e⁻s (filled)

Screening power

$s > p > d > f$

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Objective of today's class

Periodic Trends: Ionic Radius & Ionisation Enthalpy

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4. Ionic Radius:

↳ It is the effective distance from the centre of the nucleus to the outermost e^- in an ion.

Ionic Radius

→ Cationic Radius

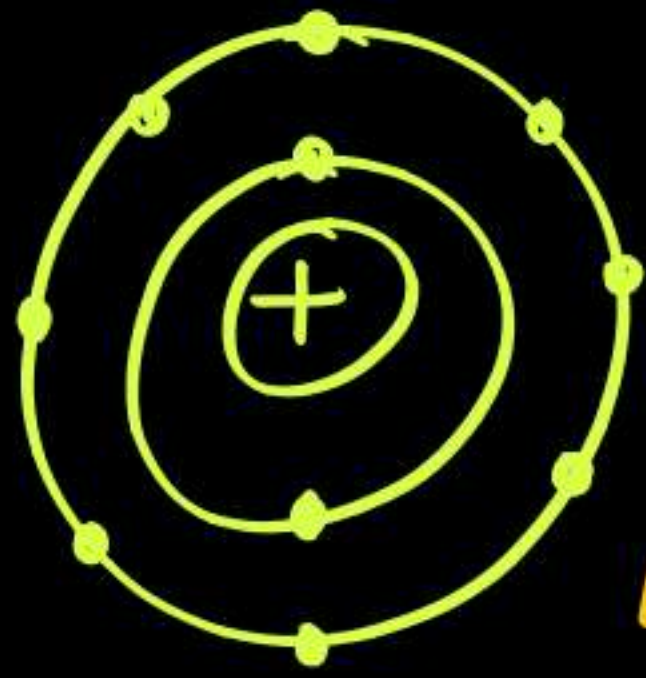
→ Anionic radius

(I.) Cationic Radii :

Radius : $_{11}Na > {}_{11}Na^{+}$

$p = 11$, $e = 11$; $p/e = 1$

$p = 11$, $e = 10$; $p/e = 1.1$



$\boxed{p/e \propto \frac{1}{\text{size}}}$

$\boxed{r_{Na} > r_{Na^{+}}}$

or $\boxed{r_{\substack{\text{neutral}\\ \text{atom}}} > r_{\text{cation}}}$

Order of radii :

(1) $_{11}Na^{+} > {}_{12}Mg^{2+} > {}_{13}Al^{3+}$

$p/e = 11/10 = 1.1$ $\quad 12/10 = 1.2 \quad 13/10 = 1.3$

(2) $Mn^{2+} > Mn^{+7}$

(3) $Fe^{3+} < Fe^{2+}$

$p/e \downarrow$, size $\uparrow$

{ Isoelectronic species → same no. of e^-s }

A.G. Mam trick

$\boxed{\text{+ve charge} \uparrow, \text{ size} \downarrow}$

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(ii.) Anionic Radii:

$$_9F < {}_9F^-$$

$p = 9$, $e = 9$, $p/e = 1$ $\qquad$ $p = 9$, $e = 10$, $p/e = 0.9$

$$\boxed{\downarrow p/e \propto \frac{1}{\text{size}\uparrow}}$$

$$\boxed{r_{\text{neutral atom}} < r_{\text{anion}}}$$

e.g. Size: $_7N^{3-} > {}_8O^{2-} > {}_9F^-$

	$_7N^{3-}$	$_8O^{2-}$	$_9F^-$
p	7	8	9
e	10	10	10
p/e	(0.7)	(0.8)	(0.9)

(Iso e^- species)

(2) $S^- < S^{2-}$

☆ A.G. Mam: trick $\boxed{-ve\ charge \uparrow \quad size \uparrow}$

$\left\{ size \propto \frac{-ve\ charge}{+ve\ charge} \right\}$

$$\boxed{r_{anion} > r_{\substack{neutral \\ atom}} > r_{cation}}$$

Trend: $Na^+ < K^+ < Rb^+ < Cs^+$

Q $N^{3-}, Mg^{2+}, Al^{3+}, O^{2-}, F^-, Na^+, Si^{4+}$

Inc. order: $Si^{4+} < Al^{3+} < Mg^{2+} < Na^+ < F^- < O^{2-} < N^{3-}$

Questions

Q1. Which has larger size?

(i) K^+ or K ✓

(ii) O^- or O^{2-} ✓

(iii) ✓ Sn^{2+} or Sn^{4+}

(iv) P^{5+} or P^{3+} ✓

(v) ✓ Mn^{2+} or Mn^{4+}

$\left\{ \uparrow \text{size} \propto \dfrac{-ve}{\downarrow +ve} \right\}$

Q2. Arrange in <u>increasing order of</u> size:

(i) I, I^+, I^- : $I^+ < I < I^-$

(ii.) $Mg, Mg^{2+}, Mg^{+}, Mg^{2-}$: $Mg^{2+} < Mg^{+} < Mg < Mg^{2-}$

(iii.) H, F, C, Na, S : $H < F < C < S < Na$

(iv.) Ni, Cu, Zn : $Ni < Cu < Zn$

(v.) Li, Na, Be : $Be < Li < Na$

$\left\{ \begin{matrix} 2^{nd} & Li & Be \\ 3^{rd} & Na & \end{matrix} \right\}$

(vi.) Ti, Zr, Hf : $Ti < Zr \approx Hf$

(vii.) Rb, Na, Cl : $Cl < Na < Rb$

Na ——— - - - Cl

K

Rb

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Ionisation Energy

- **Ionisation Energy:** (I.E.) It is the minimum amount of energy required to remove the outermost shell e^- in an isolated gaseous atom.

  isolated → (non-bonded)

$$X(g) \xrightarrow{I.E.} X^+(g) + e^-$$

$$\boxed{\text{Unit} = KJ/mol}$$

{ Ionisation energy = Endothermic (+ve) }

- **Ionisation Potential:** (I.P.) It is the amount of potential difference required to remove an e^- from its outermost shell in an isolated gaseous atom.

$$\boxed{\text{Unit} = eV/atom \text{ or } eV/mole}$$

$\{1 eV = 1.6 \times 10^{-19} J\}$

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$X(s) \longrightarrow$ removal $e^- \longrightarrow$ Not possible.

$$X(s) \longrightarrow X(g)$$ [$\Delta H_{sublimation}$: endo (+ve)]

$$X(g) \longrightarrow X^+(g) + e^-$$ [$\Delta H_{ionisation}$: +ve (endo)]

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❖ Successive Ionization Energy:

Step1: $X(g) \xrightarrow{I.E_1} X^+(g) + e^-$

Step2: $X^+(g) \xrightarrow{I.E_2} X^{2+}(g) + e^-$

Step3: $X^{2+}(g) \xrightarrow{I.E_3} X^{3+}(g) + e^-$

$I.E_1 = 1^{st}$ ionization energy

$I.E_2 = 2^{nd}$ step I.E

$I.E_3 = 3^{rd}$ step I.E.

Successive I.E: $\boxed{I.E_1 < I.E_2 < I.E_3}$ {for a given element}

e.g. $_{13}Al \xrightarrow{I.E_1} {}_{13}Al^+ \xrightarrow{I.E_2} {}_{13}Al^{2+} \xrightarrow{I.E_3} Al^{3+}$

$_{13}Al$	$_{13}Al^+$	$_{13}Al^{2+}$	Al^{3+}
p=13	p=13	p=13	p=13
e=13	e=12	e=11	e=10

$IE_1 < IE_2 < IE_3$

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❖ Factors affecting the ionisation energy :

(i) Atomic size : Ionisation energy is inversely proportional to the size of an atom. As size increases, force of attraction on outer most electron decreases and hence electron can be removed easily.

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

(ii) Nuclear charge : The ionisation energy increases with increase in effective nuclear charge

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

(iii) Penetration effect of the electron : Different subshells have different penetration power or ability to come closer to the nucleus. Penetration power of a subshell is in the order $\boxed{s > p > d > f}$ [within same energy level]. Greater will be the penetration power of a subshell, closer will be the electron to the nucleus, and higher will be its ionisation energy.

$$\boxed{\text{Penetration Power} \uparrow, \ I.E. \uparrow}$$

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(iv) Shielding effect of the inner shell electron : As the shielding effect increases, the ionisation energy decreases.

Shieding effect ↑ ; I.E ↓

Imp.

(v) Stability of half-filled & fully-filled orbitals:

↳ $p^3/d^5/f^7$ ↳ $p^6/d^{10}/f^{14}$ (Extra-Stable)

e.g. ${}_7N$ $(1s^2 2s^2 2p^3)$ ${}_8O$ $(1s^2 2s^2 2p^4)$

$\boxed{\uparrow\downarrow}$ $\boxed{\uparrow \mid \uparrow \mid \uparrow}$ $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow \mid \uparrow \mid \uparrow}$

2s 2p³ 2s 2p⁴

I.E$_1$ $2p^3 \rightarrow 2p^2$ > $2p^4 \rightarrow 2p^3$

I.E$_2$ <

I.E$_1$	N > O
I.E$_2$	O > N

{After removing 1st e⁻; 'O' attains the E.C. of N}

$_4Be$

$1s^2\ 2s^2$

$\boxed{\uparrow\downarrow}$

$_5B$

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

$\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\ |\ \ |\ \ }$

$I.E_1$	Be	$>$	B
$I.E_2$	Be	$<$	B

{ Penetration power
$s > p > d > f$ }

Be
$2s^2$
$\downarrow$
$2s^1$

B
$2s^2 2p^1$
$\downarrow$
$2s^2$

❖ Trend of ionization energy in group & period:

In a period, at. size ↓ses & Z_{eff} ↑ses on moving in a period,

I.E ↑ses (generally)

2^{nd} period ($I.E_1$): $Li < \boxed{Be > B} < C < \boxed{N > O} < F < Ne$

Be $(2s^2)$, B $(2s^2 2p^1)$, N $(2p^3)$, O $(2p^4)$

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

3^{rd} period ($I.E_1$): $Na < \boxed{Mg > Al} < Si < \boxed{P > S} < Cl < Ar$

Mg $(3s^2)$, Al $(3s^2 3p^1)$, P $(3p^3)$, S $(3p^4)$

In a group, size ↑se & I.E. ↓se

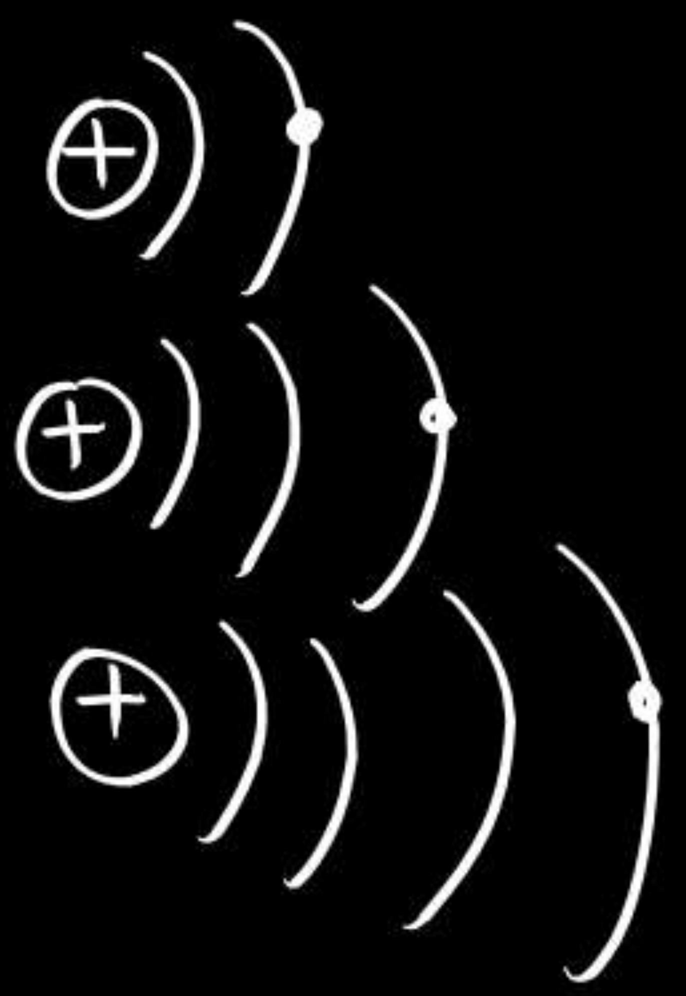
Li

Na

K

Rb

Cs



I.E ↓se

PERIOD →

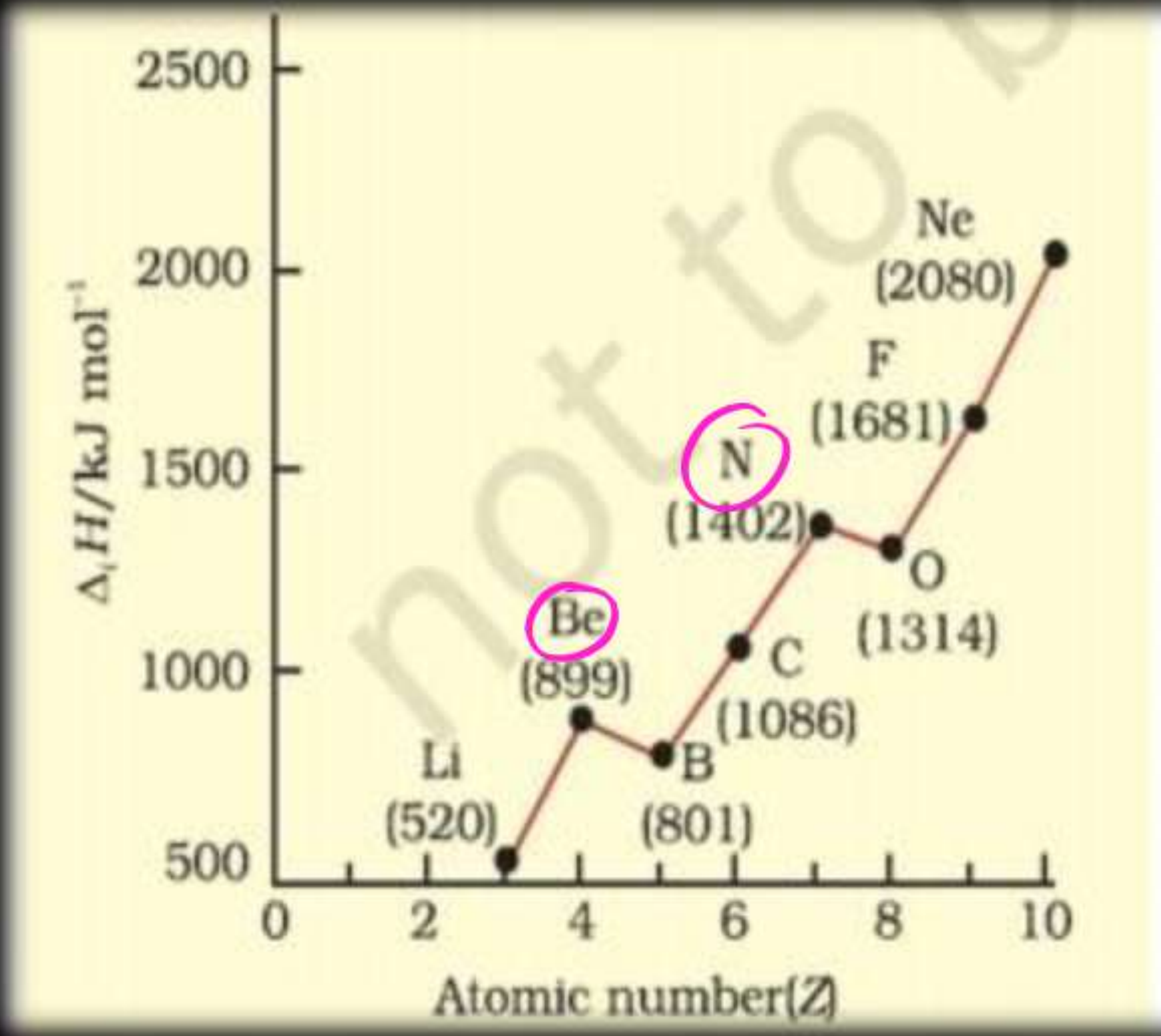
Size ↓

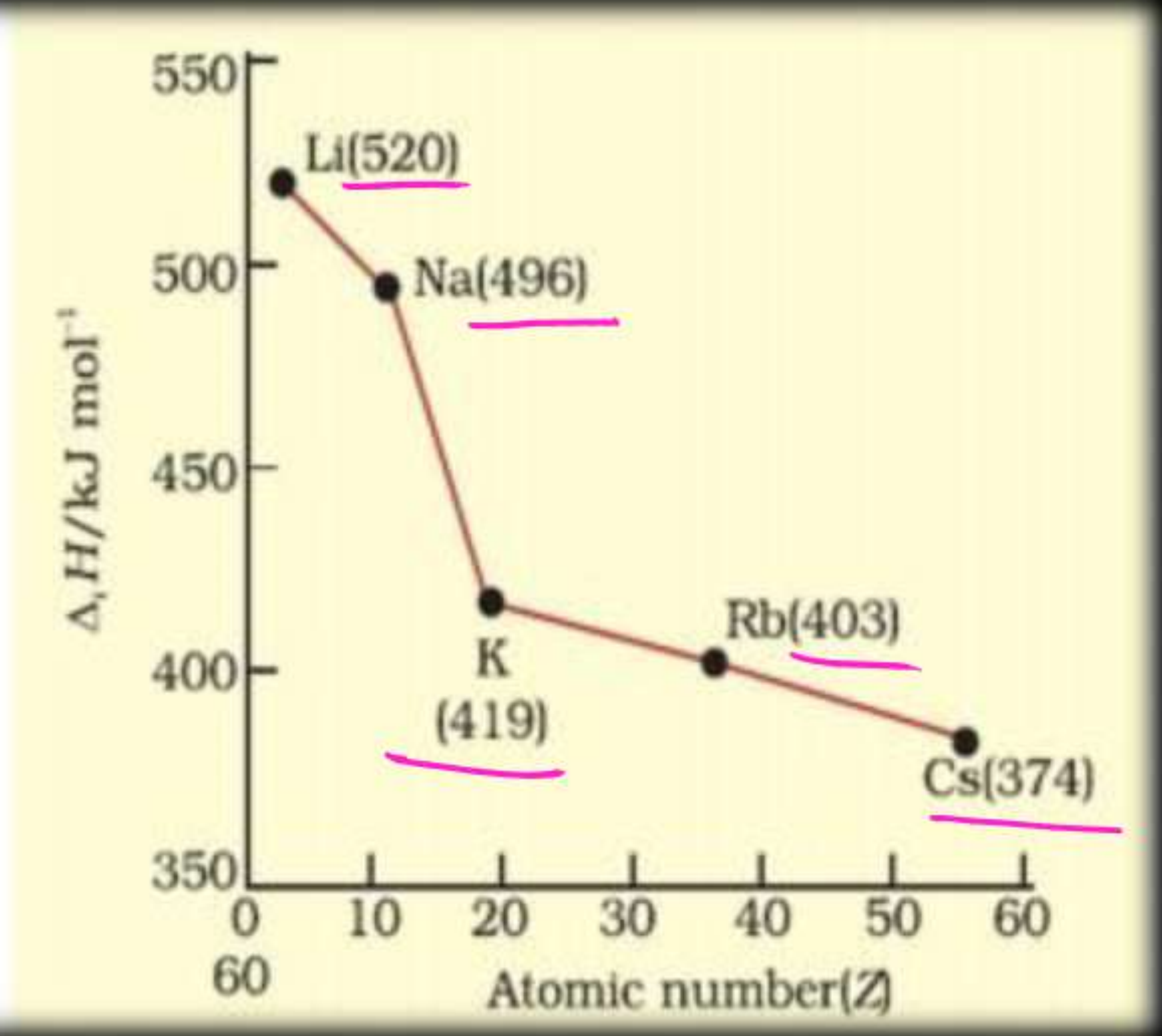
I.E ↑

GROUP ↓

Size ↑

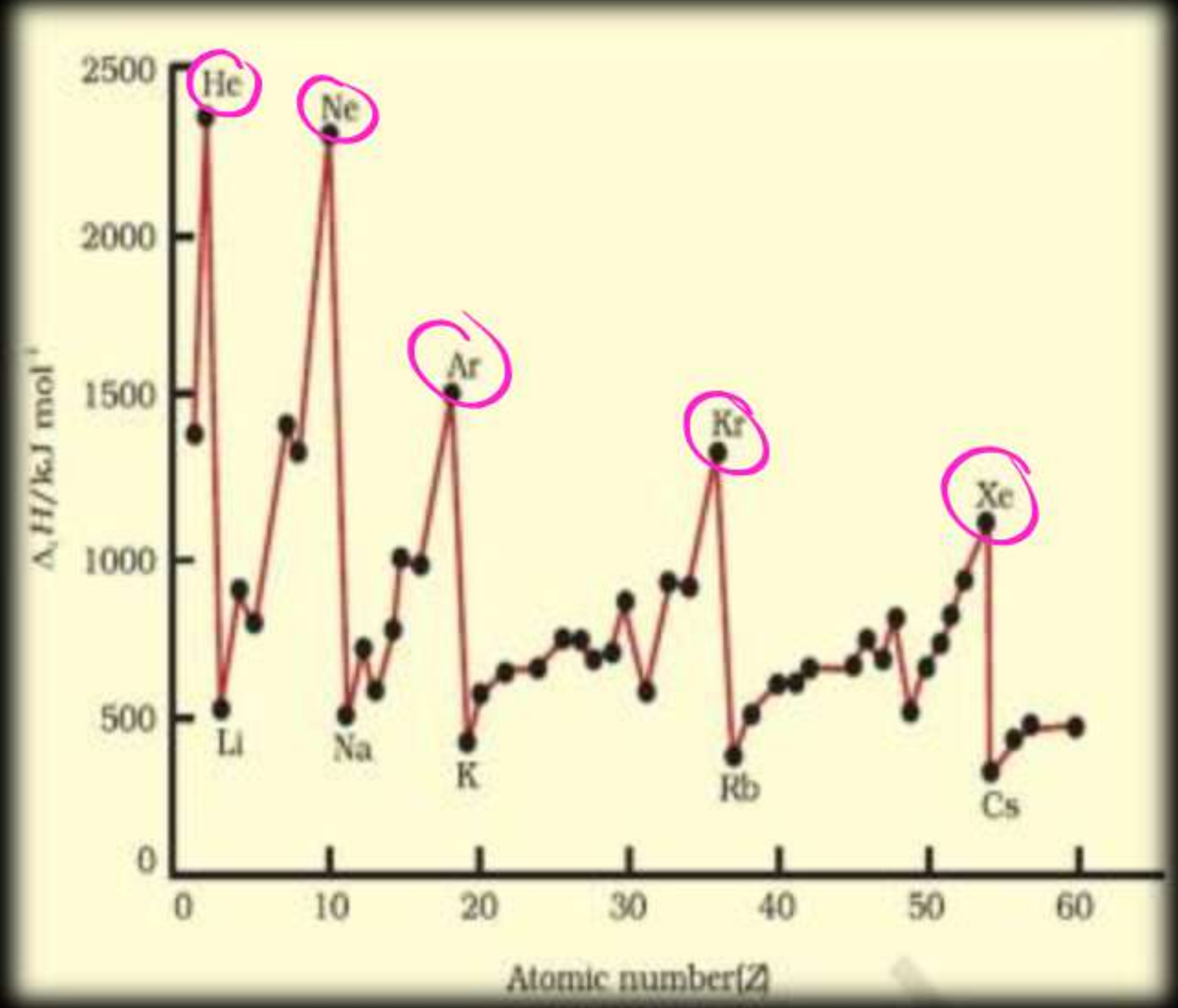
I.E ↓





2nd Period (I.E$_1$)

Gp 1 (I.E$_1$)



Exceptions:

1. I.E. of Ga > Al

$$\boxed{I.E \propto \frac{1}{size}}$$

(size): Ga < Al

($I.E_1$) Ga > Al

2. I.E. of 5d > 4d

$3d < 4d \approx 5d$

I.E $\boxed{Zr < Hf}$

size: (160 pm) (159 pm)

APPLICATION OF I.E.

1. Metallic & Non-metallic character:

$$\boxed{\downarrow I.E \propto \frac{1}{\text{Metallic char}}}\uparrow$$

$\left\{ \begin{array}{l} \text{Metals} \longrightarrow \text{to lose } e^- \longrightarrow \text{low I.E} \\ \text{Non-Metals} \longrightarrow \text{high I.E.} \end{array} \right\}$

2. Reactivity of Metals:

→ ease of losing e^-

$$\boxed{\downarrow I.E \propto \frac{1}{\text{Reactivity of metals}}}\uparrow$$

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3. To determine the no. of valence electrons of an element:

$$\boxed{\text{No. of valence electrons} = \text{no. of lower values of } \underline{\text{I.P.}} \text{ before } \underline{1^{st} \text{ highest}} \text{ jump}}$$

4. Stability of oxidation states of an element:

$\boxed{\Delta \text{I.E.} \geq 16\ \text{eV}}$ [Lower Oxd^n State : Stable]

$\boxed{\Delta \text{I.E.} \leq 11\ \text{eV}}$ [Higher Oxd^n State : Stable]

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Questions

Q.1 I.P$_1$ and I.P$_2$ of Mg are 178 and 348 Kcal mol^{-1}. The enthalpy required for the reaction: $Mg \longrightarrow Mg^{2+} + 2e^-$ is

(i) +170 Kcal
(ii) +526 Kcal
(iii) −170 Kcal
(iv) −526 Kcal

Sol:

$$Mg \longrightarrow \cancel{Mg^+} + e^- \qquad IP_1 = 178 \text{ Kcal/mol}$$
$$+ \quad \cancel{Mg^+} \longrightarrow Mg^{2+} + e^- \qquad + \; IP_2 = 348 \text{ Kcal/mol}$$
$$Mg \longrightarrow Mg^{2+} + 2e^- \qquad I.P. = +526 \text{ Kcal/mol}$$

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Q2. The $I.P_1$, $I.P_2$, $I.P_3$, $I.P_4$ and $I.P_5$ of an element are 7.1, 14.3, 34.5, 46.8 & 162.2 eV respectively. The element is likely to be

(i) Na : 2, 8, 1 (ii) Si : 2, 8, 4 (iii) F : 2, 7 (iv) Ca : 2, 8, 8, 2

Sol: No. of valence e^-s : lower IP values before 1^{st} highest jump

= 4

Q3. $M(g) \longrightarrow M^+(g) + e^-$ $\quad I.E_1 = 7.9 eV$

$M^+(g) \longrightarrow M^{2+}(g) + e^-$ $\quad I.E_2 = 15.5 eV$

Which oxidation state is more stable?

(i) M^+ $\quad$ (ii) M^{2+} $\quad$ (iii) Both $\quad$ (iv) None

Sol:
$$\Delta I.E = I.E_2 - I.E_1$$
$$= 15.5 - 7.9$$
$$= 7.6\ eV$$

$\Delta I.E \leq 11 eV$ (Higher O.S. $\longrightarrow$ stable)

Homework

1. Read NCERT from Pg. No. → 74 to 89
2. Solve DPP-4 & 5
3. Revise all the notes from starting.

Thank You