UGTRB CHEMISTRY

UNIT – 1

PERIODIC PROPERTIES

CLASS- 4

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Ionic radii :

- Distance between closest positive and negative ions in the lattice Crystal

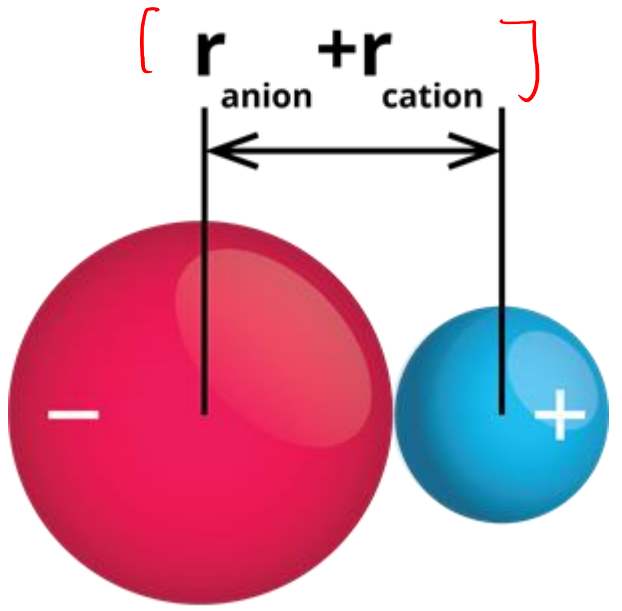
- ✓ Determined by x - ray diffraction

Ionic Compd = Solid

✓ General trend : atomic = Ionic

L to R ⇒ decrease

T to B → Increase



ionic radius

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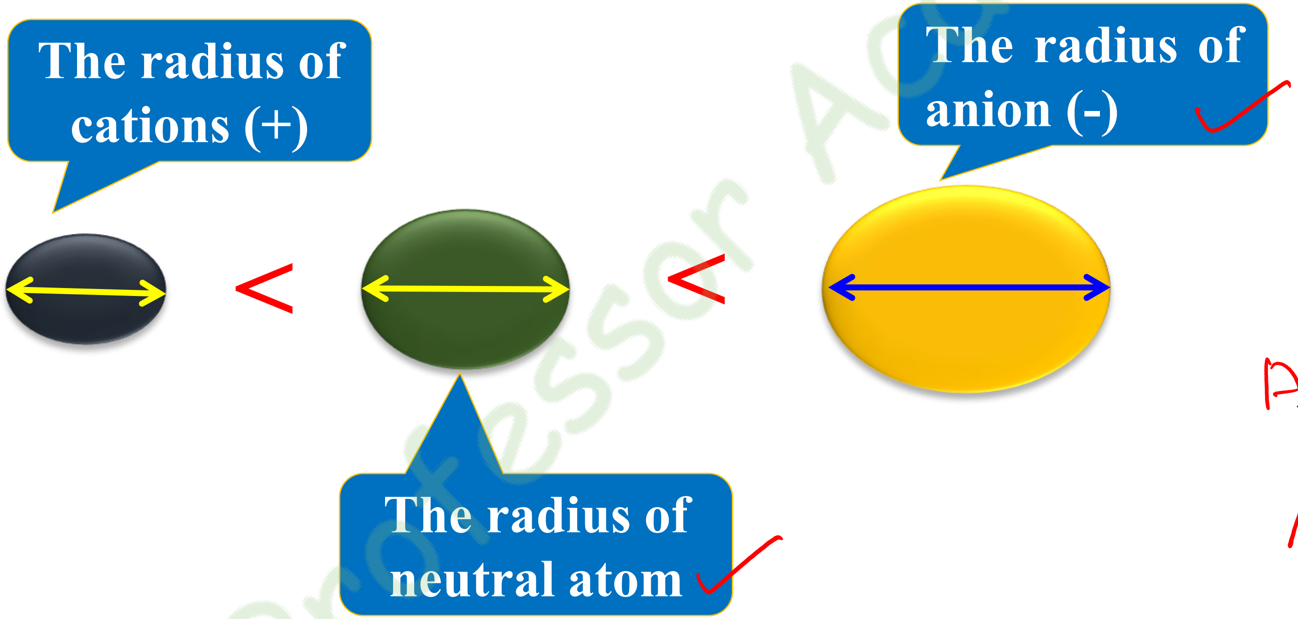
Variation in radius of an element :



The radius of cations (+) < The radius of neutral atom ✓ < The radius of anion (-) ✓

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Ionic radii

↓

ions

Cation

anion's

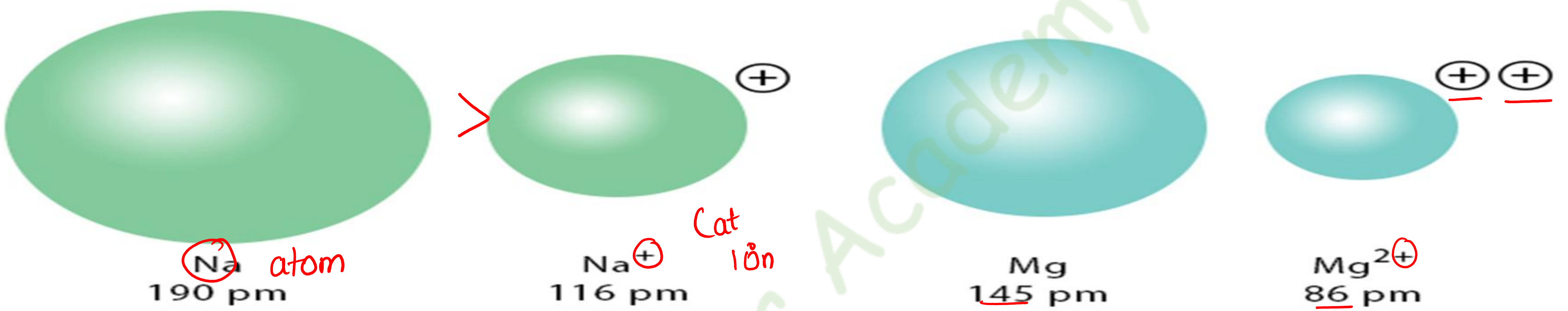
$\text{Atom} + e^- \rightarrow \text{anion}$

$\text{Atom} - e^- \rightarrow \text{cation}$

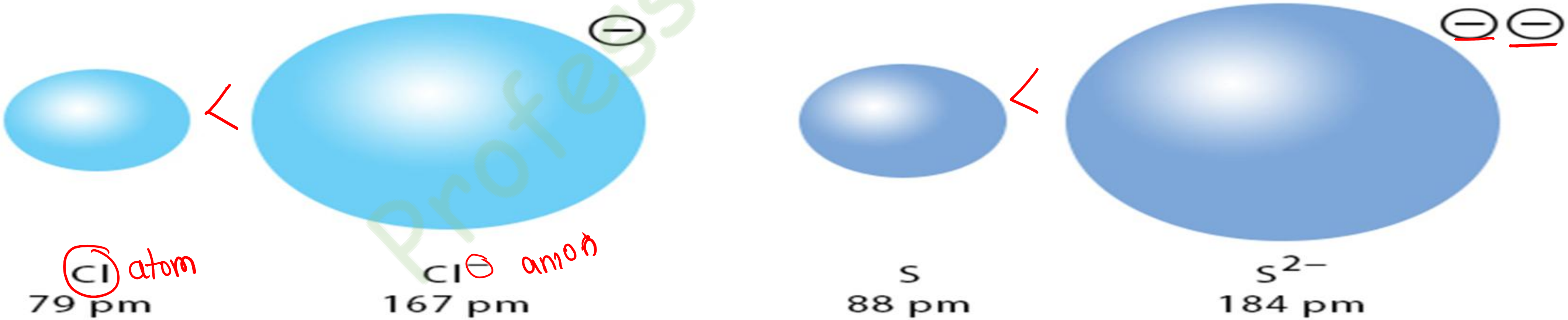
Ionic & atom ↷ $\boxed{\text{Cation} < \text{atom} < \text{anion}}$

Atomic Radius vs Ionic Radius

Cation: Atomic Radius > Ionic Radius



Anion: Atomic Radius < Ionic Radius



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more +ve charge on atom ⇒ Small size

❖ cation radius < atomic radius

more -ve charge ⇒ Large size

❖ Anion radius > atomic radius

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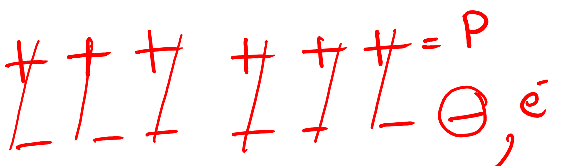
Ex:

Carbon atom p = 6
e = 6

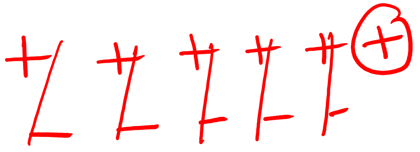
\+ + + + + +
\- - - - - - } Neutral

$C^{\ominus}$: p = 6
e = 7



Repulsion high
Size high

C^{+} ⇒ p = 6
e = 5



Proton high
attraction high
∴ Size low.

The Correct order of the size of S, S^{-2}, S^{+2} and S^{+4} species is

atom anion cation

1) $S > S^{+2} > S^{+4} > S^{-2}$

2) $S^{+2} > S^{+4} > S^{-2} > S$

anion atom cation small size

3) $\underset{\text{anion}}{S^{-2}} > \underset{\text{atom}}{\underline{S}} > \underset{\text{cation}}{S^{+2}} > \underset{\text{small size}}{S^{+4}}$

4) $S^{+4} > S^{-2} > S > S^{-2}$

anion > atom > cation

$\boxed{\text{AAC}}$

$\boxed{\underset{-ve}{\text{high}} > \underset{-ve}{\text{low}} > \text{atom} > \underset{+ve}{\text{low}} > \underset{+ve}{\text{high}}}$

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Isoelectronic Species

Same e^-

- Atoms, ions, or molecules that have the **same total number of electrons** and thus share the **same electron configuration**, even though they might be different elements with varying nuclear charges

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No Isoelectronic Species
↓
General trend"

10 = 8+2 ↑

10 = 9+1

11−1 = 10

12−2 = 10

- For example, O^{2-}, F^{-}, Na^{+}, and Mg^{2+} all have 10 electrons, making them isoelectronic with Neon (Ne).

10

10

- This similarity in electron count means they often exhibit similar chemical behaviors and properties

For Isoelectronic Species, Ionic radii Based on Nuclear charge.

$\boxed{\Rightarrow O^{-2} > F^{-} > Ne > Na^{+} > Mg^{+2}}$ ★

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" Ionisation energy/

Ionisation enthalpy/

Ionisation Potential "

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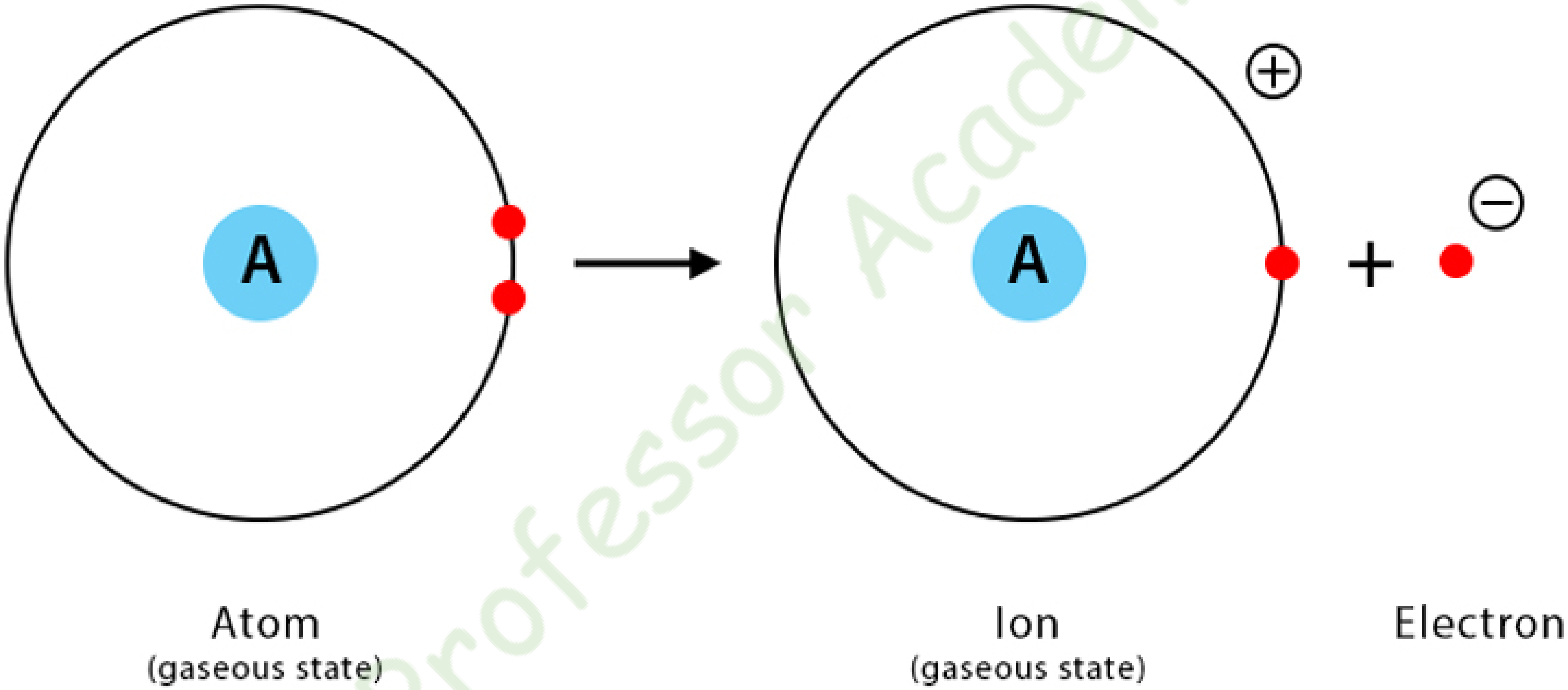
Ionization Energy



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- (The minimum energy required to remove an electron from an isolated gaseous atom)

$$A^{n+}(g) \longrightarrow A^{(n+1)+}(g) + \boxed{e^{-}} \qquad \text{ionization energy } (IE) = \Delta U$$

cation

- Expressed in unit - "kJ mol–1 or kcal/mole or ev/atom"

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energy absorb = endo ⊕
energy release = exo ⊖

- Energy is always required to remove electrons from an atom (endothermic) process

- Ionization enthalpies are always positive

$\boxed{IE = +ve}$

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NOTE

$IE_2 > IE_1$, Because It Is More Difficult To Remove An Electron From A Positively Charged Ion Than From A Neutral Atom. So more energy required

General trend :

L to R ⇒ ↑ses (Z^*eff)

T to B ⇒ ↓ses (shielding effect)

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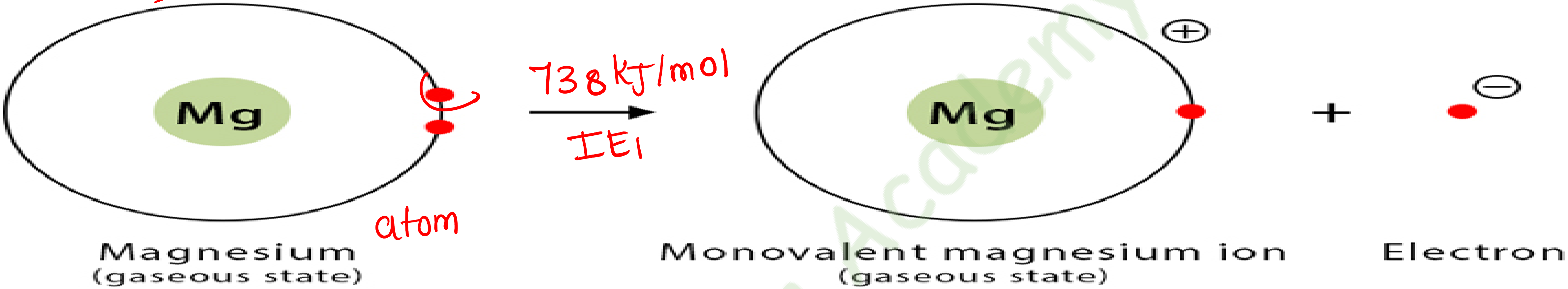
Second Ionization Energy

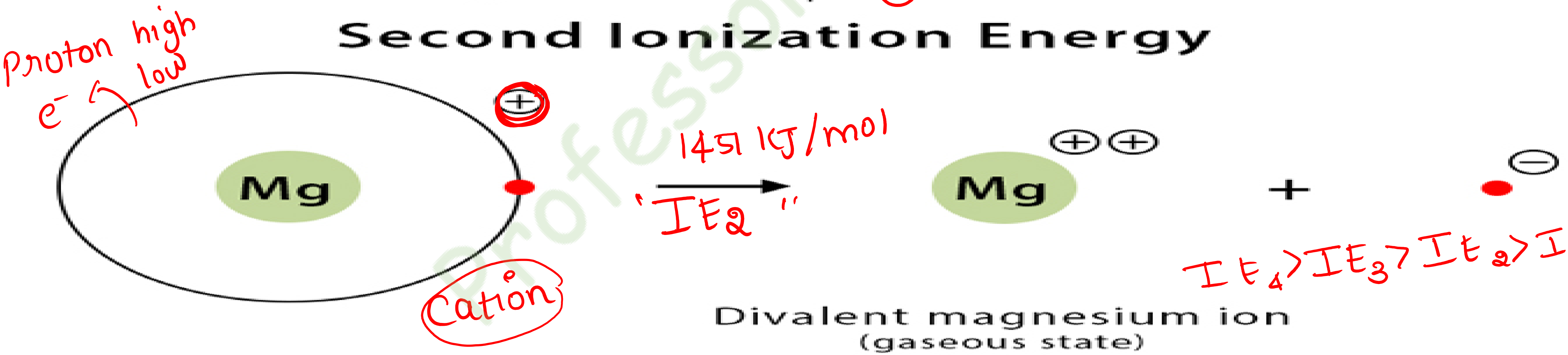
p = e

First Ionization Energy



Energy absorbed: $I_1 = +738$ kJ/mol

Second Ionization Energy



$IE_4 > IE_3 > IE_2 > IE_1$

Energy absorbed: $I_2 = +1451$ kJ/mol

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Factors Affecting The Ionization Potential

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Factors affecting the Ionization Potential

1) Atomic Size :

- When atomic size increases, I.P. value decreases because the attraction of nucleus over valence electrons decreases

$$\text{i.e.,} \quad \text{I.P.} \propto \frac{1}{\text{Atomic size}}$$

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2) Nuclear Charge

- **When nuclear charge increases, I.P value also increases because the attraction of nucleus over the valence electron increases.**

i.e., I.P. $\propto$ Nuclear charge

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3) Screening effect

- When screening effect increases, I.P value decreases because the attraction of nucleus over valence e^- decreases.

$$\text{i.e.,} \quad \text{I.P.} \propto \frac{1}{\text{Screening effect}}$$

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4) Penetrating effect

What is penetrating effect?

- The order of penetrating power of various orbitals is…

  s > p > d > f

- When penetrating effect increases, I.P. value increases because attraction of nucleus over the electrons increases.

i.e., I.E. ∝ Penetrating effect

The tendency of orbital to move towards the nucleus is called penetrating effect

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5) Electronic configuration

- The atoms of elements having either half filled or fully filled orbitals are comparatively more stable hence they requires more energy than others.

Ex: I.E. of N > I.E. of O

$p = 6e^-$

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

2p

| ↑ | ↑ | ↑ |

half filled

>

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

2p

| ↑↓ | ↑ | ↑ |

more than half filled

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$$I.E \propto \frac{Z^{*} \propto \text{Pentrating}}{\text{size, Screening constant}} \propto \text{half \& fully configurar.}$$

:: Make a note ::

Group 18 elements have maximum I.E. in a period due to completely stable electronic configuration i.e. octet configuration(ns^2np^6)

I^{st} Period H He ✓

$$\boxed{H < He}$$

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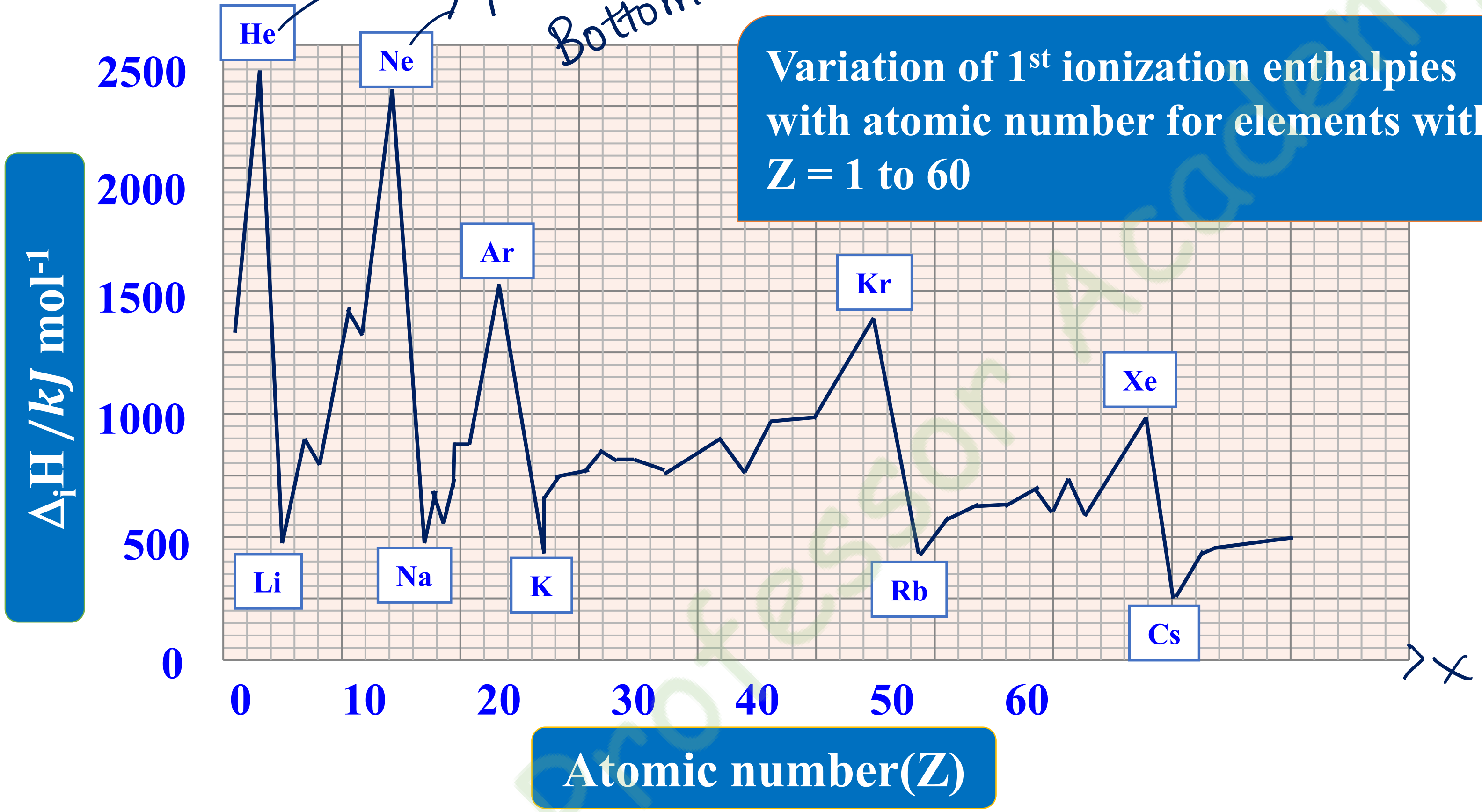
Variation of 1st ionization enthalpies with atomic number for elements with Z = 1 to 60

peak = Noble gas

Bottom = alkali metal



Stable

I.E high = Noble gas

I.E low = alkali metal

Reactive

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	Li	Be	B	C	N	O	F
Configuration	$2s^1$	$2s^2$	$2s^22p^1$	$2s^22p^2$	$2s^22p^3$	$2s^22p^4$	$2s^22p^5$
Type of orbital	s	s	p	p	p	p	p
Radii (Å)	1.34	1.25	0.90	0.77	0.74	0.73	0.71
Z*	1.30	1.95	2.6	3.25	3.90	4.55	5.20
IE (ev)	5.39	9.32	8.29	11.26	14.53	13.62	17.42

Penetrating

2nd period I.E order : $Li < B < Be < C < O < N < F$

$Z=5$ $1s^2 2s^2 2p^1$

$Z=4$ $1s^2 \boxed{2s^2}$

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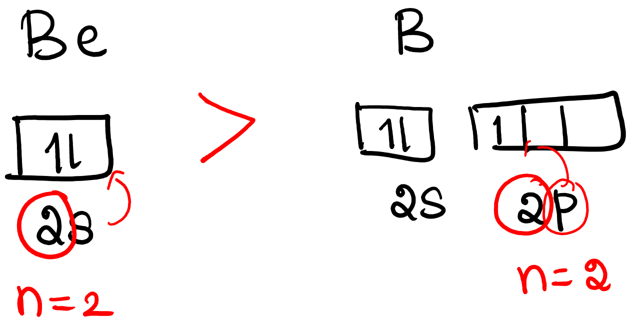
Shell value same, diff orbital
↓
Penetrating



Shell
n=2

Be > B

$2s^2$

$2s^2\ 2p^1$

Penetration effect

(a) **First IE of Be is greater than the first IE of B**

The reason is (difference in the nature of valence orbitals.) The valence orbitals of Be is $2s^2$ and that of B is $2s^2 2p^1$. An s orbital is more penetrating and less shielded whereas a p – orbital is less penetrating and more shielded. Therefore, attractive effect is larger on s- electron than on p – electron. Beryllium, so has higher IE than B.

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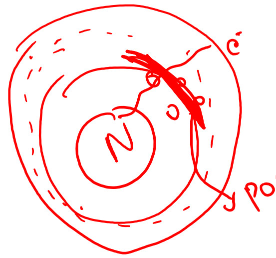

4f
poor shielding
size ↓ses

(b) **The first IE of N is greater than Oxygen**

In this case, the reason is difference in the stability of electron configuration. Nitrogen has full filled 2s orbital and half filled 2p orbitals i.e., both are "extra stable". Oxygen has full filled 2s orbital but its p – orbitals are neither half nor full filled i.e., not extra stable. Therefore, removal of p – electron of oxygen is easy and its 1st IE is lower than N.

Scandide

(c) The first IE of Al and Ga are almost equal (reason, equal radii)

I.E Ga ≃ Al

3d

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

(d) Variation of 1st IE along each transition series is small (reason, small difference in radii)

(e) The first IE of third transition series elements is larger than second transition series elements (reason – lanthanide contraction)

I.E : $\boxed{4d < 5d}$

Small I.E high

(f) The change in first IE for lanthanide elements is small (reasons, small change in the radii of lanthanide elements from Ce to Lu).

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Cr ⇒ Periphery

↙ 24

$$\underbrace{(1s^2)}_{(n-3)}\underbrace{(2s^2\ 2p^6)}_{(n-2)}\underbrace{(3s^2\ 3p^6)(3d^5)}_{\underline{(n-1)}}\underbrace{(4s^1)}_{\underline{n}}\ \xrightarrow{e^-}\ \text{Imaginary } e^-$$

$$S\,e^-/p\,e^- = (n \times 0.35) + (n-1 \times 0.85) + (n-2 \times 1) \ldots$$

$$= (1 \times 0.35) + 13 \times 0.85 + 8 \times 1 + 2 \times 1$$

$$= 0.35 + 11.05 + 8 + 2$$

$$\boxed{\sigma = 21.4}$$

$$Z^* = Z - \sigma$$

$$= 24 - 21.4 = \boxed{2.6}$$

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H.W

Z^* eff: Na

Cl^-

F^-

C

Ni

Pairing: N

S

Fe

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