Ionization Enthalpy :-

- A quantitative measure of the tendancy of an element to lose electron is given by its Ionization Enthalpy.
- It is the amount of energy required to remove an electron from an isolated gaseous atom (x) in its ground state. It is represent by IE
- The ionisation energy required to remove first electron from neutral atom is termed as first ionisation energy IE_1. Similarly energy required to remove second and third electrons are referred as second ionisation energy (IE_2) and third ionization energy (IE_3) respectively.

$$M_{(g)} \longrightarrow M^{+}_{(g)} + e^- \qquad \Delta H = IE_1$$
$$M^{+}_{(g)} \longrightarrow M^{2+}_{(g)} + e^- \qquad \Delta H = IE_2$$
$$M^{2+}_{(g)} \longrightarrow M^{3+}_{(g)} + e^- \qquad \Delta H = IE_3$$

*** $IE_1 < IE_2 < IE_3$ always.

* unit of IE is KJ/mol or Kcal/mol or ev/atom
* Ionization enthalpies are always positive
* The term "ionization enthalpy", if not qualified, is taken as the first ionization enthalpy.

factors affecting Ionization Enthalpy

- Ionisation energy value depends upon the force of attraction between nuclei and electron which has to be removed.

(i) Effective nuclear charge :→

$$IE \propto Z_{eff}$$

(ii) size of atom :→ size ↑ses I.E. ↓ses

(iii) Subshell from which electron is removed :-
Lesser is energy of sub-shell, closer it is to the nucleus ∴ order of IE $s > p > d > f$
∴ IE_1 Be > B

(iv) Nature of configuration :-
Half filled & full filled orbitals are more stable than partially filled orbital ∴ removal of e⁻s from these configuration is difficult.
∴ IE_1 of N > IE_1 of O
IE_2 of O > IE_2 of N

(v) Shielding effect or screening effect: more is the shielding of valence electrons, Lower is the IE.

Variation of IE :

- IE values of inert gases are exceptionally high due to closed electron shell and very stable electron configuration.
- Alkali Metals have very low ionisation energy & it can be correlated with their high reactivity.
- As we go across the period IE_1 generally ↑ses: when we move from Lithium to fluorine across the period, successive electrons are added to orbitals in the same principal quantum level and the shielding of the nuclear charge by the inner core electrons does not increase very much to compensate for increase in attraction of the electron to the nucleus. Thus, across a period, increasing nuclear charge outweighs the shielding. Consequently, the outermost electrons are held more and more tightly and the ionization enthalpy increases across a period. A/R
- As we go down a group IE decreases. Because as we go down a group, the outermost electron being increasingly farther from the nucleus, there is an increased shielding of the nuclear charge by the electrons in the inner levels. In this case, increase in shielding outweighs the increasing nuclear charge and the removal of the outermost electron requires less energy down a group. A/R

Exception & Reason :-

Li < B < Be < C < O < N < F < Ne (I.P. Value)

- First ionisation enthalpy of Boron (Z=5) is slightly less than that of beryllium (Z=4) even though the former has a greater nuclear charge

  Reason :- The penetration of a 2s-electron to the nucleus is more than that of a 2p-electron hence the 2p electron of boron is more shielded from the nucleus by the inner core of electrons than the 2s electrons of beryllium. Therefore, it is easier to remove the 2p-electron from boron compared to the removal of a 2s-electron from beryllium. Thus, boron has a smaller first I.E. than beryllium. A/R.

- There is smaller first ionisation ethalpy of oxygen compared to Nitrogen. A/R.

  Reason :- This occur because in nitrogen atom, there are three 2p-electrons present in different atomic orbital (Hund's Rule) whereas in the oxygen atom, two of the four 2p-electrons must occupy the same 2p-orbital resulting in an increased electron-electron repulsion. Consequently, it is easier to remove the forth 2p-electron from oxygen that it is to remove one of the three 2p-electrons from nitrogen. A/R

- Pd > Ag → reason electronic configuration
  $4d^{10}, 5s^0$ $4d^{10} 5s^1$

→ Cd > In
→ Hg > Tl
→ 13^{th} grp ⇒ Tl > In
14^{th} grp = Pb > Sn

→ He has higest I.E.
→ Cs has Lowest I.E.
→ Fr > Cs
→ Ra > Ba
→ Poor shielding of e⁻ by inner f-electrons

- If $\Delta IE_{(1,2)}$ = 965-1450 KJ/mol or less than Higher ox. state will be More stable
→ If $\Delta IE_{(1,2)}$ > 1450 KJ/mol, then lower oxi. state will be more stable.