1. IONIZATION ENERGY (IE):-

ionization energy is the energy required to remove an electron from a gaseous atom. Generally, IE increases across a period and decreases down a group. However, there are some notable exceptions.

- Group 13 (Boron family) Vs Group 2 (Alkaline Earth metals):-
  - The first ionization energy of Group 13 elements (eg: Boron) is lower then that of Group 2 elements (eg Beryllium) in the same period.
  - REASON:- Group 13 elements, have a single electron in a higher 'p' orbitals, which is slightly higher in energy and experiences more shielding from the 's' electrons, making it easier to remove.

Group 16 (Chalcogens) Vs Group 15 (Phictogens):-

- The first ionization energy of Group 16 elements (eg. Oxygen) is lower then that of Group 15 elements. (eg. Nitrogen) in the same period.
- REASON:- Group 15 elements have a half-filled 'p' subshell which provides extra stability. Removing an electron from Group 16 elements disrupts a doubly occupied 'p' orbital, leading to "electron-electron" repulsion that makes it easier to remove that electron.

2. ELECTRON AFFINIY (EA):-

Electron affinity is the energy change when an electron is added to a gaseous atom. Generally, ~~electron is added to a gaseous~~ EA becomes more negative (more favourable) across a period and less negative down a group.

- Group 17 (Halogens) Vs Group 18 (Nobel Gases):-
  - Halogens have the most negative (most favorable) electron affinityes because adding an electron completes their octet, achieving a very stable nobel gas configuration.
  - Nobel gases have positive electron affinity, means energy is required to added an electron, as their electron shells are already full and stable.
- Group 2 (Alkaline Earth metals) and Group 15 (Pnictogens)
  - These groups tend to have less negative (or even positive) electron affinities.
  - REASON:- Group 2 elements have a filled 's' subshell, and adding an electron would place it in a higher energy p orbital. Group 15 elements have a stable half-filled 'p' subshell. Adding an electron would disrupt this stability.
- 2nd period Elements (especially N, O, F) vs. Third Period Element (P, S, Cl):-
- Often, the electron affinity of second-period elements is less negativity than their third-period counterparts.
- Reason:- The smaller size of the second period-period elements leads to increased electron-electron repulsion when an extra electron is added, making the process less energetically favorable. For example, the electron affinity of chlorine is more negative than that of flourine.

3. ATOMIC RADIUS:-

Atomic radius generally decreases across a period and increases down a group.

- Transition Metals (d-block contraction):-
  - Across a period in the transition metals, the decrease in atomic radius is less pronounced than in main group elements.
  - Reasons:- The added electrons filled inner 'd' orbitals. which provide some shielding, counteracting the increased nuclear charge to some extent.
- Lanthanoide Contraction:-
  - Elements immediately following the lanthanoides (eg Hafinium and Tantalum) have atomic radii that are smaller than expected based on their position in the periodic table.
  - Reason:- The poor shielding effect of the "4f" electrons in the lanthanoides leads to a greater effective nuclear charge, pulling the outer electrons closer to the nucleus. this effect carries over to the subsequent elements.

4. Electronegativity:-

Electronegativity is the stability of atom to attract electrons in a chemical bond. Generally, it increases across a period and decreases down a group.

- Noble Gases (Group 18):-

Most Noble gases are not assigned electronegativity values as they generally do not form chemical bonds. However, Xenon and Krypton can form compounds with highly electronegative elements like Fluorine & Oxygen.

5. Valency and Oxidation State :-

- Variable Valency / Oxidation States :-
  - Transition metal :- These elements exhibit a wide range of oxidation states due to the involvement of both 'ns' and (n-1) d electrons in bonding. For example, Iron can be +2 or +3.
  - Lanthanoides and Actinides :- Also show variable oxidation states.
- Inert Pair Effect :-
  - Elements in Group 13, 14 and 15 in the lower periods (eg, Lead, Bismuth) show a tendency to exhibit an oxidation state that is 2 less than their group number. This is due to the reluctance ~~in bonding~~ of the 'ns' electrons to participate in bonding, as they are held more tightly by the nucleus. For example, Lead (Group 14) commonly forms +2 compounds in Addition to +4.