2. Classification of Elements and periodicity in properties

Synopsis :

- About 106 elements are clearly known so far.
- Elements upto z = 92 occur in nature.
- Elements from z = 93 on wards are artificially prepared by artificial transmutation. Therefore, these are called transuranic elements.
- First synthetic element is Technetium (Tc)
- Out of 106 elements, 83 are metals, 11 are gaseous elements, 6 are non-metallic solids, 5 are metalloids, 1 liquid non-metal. Out of 83 metals, only liquid metal is Hg at room temperature other liquid metals, little above room temperature are Ga, Cs and Fr.
- The most abundant element in earth's crust is O_2.
- The most abundant metal in earth's crust is AI
- Order of abundance of elements in earth's crust: O_2 > Si > AI > Fe > Ca > Na > K
- The most abundant element in human body by weight is O_2.
- The most abundant element in human body in terms of moles is H_2.
- The element with highest M.P. is carbon.
- The metal with highest M.P. is tungsten.
- The metal with highest B.P. Rhenium.
- The element with lowest M.P. and B.P. is Helium.
- To simplify the study of chemistry of elements and their compounds systematic classification is necessary.
- The first attempt was made by Dobereiner.
- According to Dobereiner's law of traids, every 3 elements with similar properties is considered as one triad.
- Dobereiner's triads are (a) Li, Na, K (b) Ca, Sr, Ba (c) S, Se, Te (d) CI, Br, I
- The atomic weight of the middle element is the average of the two neighbouring elements of a given triad.
- Lothermeyer is popular for his atomic volumes curve.
- Boron has least atomic volume
- Francium has largest atomic volume.
- In the plot of atomic volume Vs atomic weights of elements, the elements with similar properties will occur in similar region.

MENDELEEV'S PERIODIC LAW:

- The physical and chemical properties of elements are the periodic functions of their atomic weights.
- In the Mendeleef periodic table elements are arranged in the increasing order of their atomic weights.
- It consists 7 horizontal rows or 7 periods and 8 vertical columns or 8 groups. Each group is sub divided into two sub groups A and B.
- Mendeleef predicted properties of certain missing elements e.g. Eka boron (Scandium), Eka aluminium (Gallium), Eka silicon (Germanium)
- The atomic weights of some elements like Be, In were corrected.

- He introduced the concept of valency and observed that valency of elements belonging to a group will be equal to it's group numbers.
- The atomic weight order was reversed in four pairs of elements, which are called anomalous pairs.
  1) ^{40}Ar, ^{39}K 2) ^{59}CO, ^{58}Ni
  3) ^{128}Te, ^{127}I 4) ^{232}Th, ^{231}Pa
- 14 elements having different atomic weights which are called rare earths are placed only in one group.
- Modern periodic table was constructed on the basis of atomic numbers.
- Properties of elements also depend on the electronic configuration.

LONG FORM OF PERIODIC TABLE:

- It is based on modern periodic law or Mosley's periodic law. It states that "the physical and chemical properties of elements are the periodic functions of their atomic numbers or electronic configuration.
- The properties are repeated after regular intervals of time when the elements are arranged in the increasing order of their atomic numbers.
- Neil's Bohr constructed the modern periodic table based on the electronic configuration of the elements.
- It is a graphical representation of Aufbau principle.
- The vertical columns are called groups and the horizontal rows are called periods.
- There are altogether 18 groups and 7 periods in the long form (or) extended form of periodic table.
- Elements are arranged in the increasing order of their atomic numbers.
- From left to right the atomic number increases by one unit.
- The electron which differentiates an element from the preceding element is called the differentiating electron.
- The differentiating electron is the last coming electron of that element.
- In each period, in the first element the differentiating electron enters into s-orbital and in the last element the differentiating electron enters into p-orbital.
- The last element of the period completes the octet by attaining the stable electronic configuration $ns^2\ np^6$.
- Thus every period start with the filling of valence s-orbital and ends with the complete filling of s and p-orbitals of valence shell.
- First period contains only two elements H($1s^1$) and He($1s^2$) and it is called very short period.
- Second period contains 8 elements and it is called 1^{st} short period.
- Third period also contains 8 elements and it is called 2^{nd} short period.
- The first 3 periods are discontinuous periods.
- 4^{th} period contains 18 elements and it is called 1^{st} long period.
- 5^{th} period also contains 18 elements and it is called 2^{nd} long period.
- Sixth period is the longest period containing 32 elements.
- Elements do not exhibit horizontal similarities as they differ in the configuration.
- Some periods are broken and some periods are extended to accommodate transition elements.
- 14 elements each of 6^{th} and 7^{th} periods have been separately placed at the bottom of table to maintain uniformity and effectiveness.
- 2^{nd} period elements are Bridge elements due to their diagonal relationship.

- 3^{rd} period elements are called typical elements as they represent the properties of below elements in the respective groups.

GROUPS :

- There are 18 groups.
- They are designated as group A and Group B except VIII and '0' groups.
- VIII group consists of 3 vertical rows or 3 groups.
- '0' group consists of Noble gases.
- Groups 'A' consists of representative elements and groups 'B' and VIII group consists of transition elements.
- Elements belonging to group will exhibit similar properties due to similar valence shell configuration.
- The elements which exhibit both vertical and horizontal similarities are transition elements.
- The number of electrons in valence shell is equal to the group number.
- The seventh period is incomplete and has about 20 elements.

Period	First element	Electronic configuration	Last element	Electronic Configuration
1	H	$1s^1$	He	$1s^2$
2	Li	[He] $2s^1$	Ne	[He] $2s^2\ 2p^6$
3	Na	[Ne] $3s^1$	Ar	[Ne] $3s^2\ 3p^6$
4	K	[Ar] $4s^1$	Kr	[Ar] $3d^{10}\ 4s^2\ 4p^6$
5	Rb	[Kr] $5s^1$	Xe	[Kr] $4d^{10}\ 5s^2\ 5p^6$
6	Cs	[Xe] $6s^1$	Rn	[Xe] $4f^{14}\ 5d^{10}\ 6s^2\ 6p^6$
7	Fr	[Rn] $7s^1$	–	–

CLASSIFICATION OF ELEMENTS INTO 4 BLOCKS:

- The elements are classified into four blocks as s-block, p-block, d-block and f-block based on the orbital into which differentiating electron enters.
- This classification is based on electronic configuration.
- s-block contains 2 groups, p-block contains 6 groups, d-block contains 10 groups and f-block contains 14 groups.
- s-block is at the extreme left and p-block is at the extreme right of the periodic table.
- d-block is kept in between s-block and p-block.
- f-block is separately placed below the main body of the table.

s-block :

- Differentiating electron enters into s-orbitals of valence shell.
- It consists of I -A and II - A groups namely alkali and alkaline earth metals.
- They are
  1) Most reactive metals.
  2) Most electropositive metals.

3) Strongly reducing in nature.
4) Strong tendency to lose electrons.

- The S-block element placed in P-block is He.
- They exhibit only positive oxidation states.

p-block:

- Differentiating electron enters into p-orbitals of valence shell.
- It consists of III-A to VII-A and '0' group.
- Electronic configuration is $ns^2\ np^1$ to $ns^2\ np^5$ and $ns^2\ np^6$.
- It consists of all types of elements i.e. metals, non-metals, metalloids.
- These are more electronegative than s-block elements.
- Most electronegative elements are present in this block.
- They are also reactive elements except 'O' group.
- They exhibit positive and negative oxidation states.

d – block :

- Differentiating electron enters into d-orbitals of (n – 1) shell.
- It consists of all groups – B and VIII group. (total ten groups)
- Electronic configuration: ns^2, $(n-1)d^{1-10}$
- All d-block elements are metals.
- These are placed in 4^{th}, 5^{th}, 6^{th} and 7^{th} periods.
- They are hard metals with high M.P.'s and B.P's.

f – block :

- Differentiating electron enters into f-orbitals of (n–2) shell.
- Electronic configuration: ns^2, $(n-1)d^{(0\text{ or }1)}$, $(n-2)f^{1\text{-}14}$.
- They are present in two horizontal rows at the bottom of periodic table namely lanthanides and actinides.
- They belong to 6^{th} and 7^{th} periods.
- They belong to III-B group.
- Many of them are artificially prepared and do not occur in nature.
- They are all metals.

Classification of elements into 4 types:

- This is based on properties of elements.
- Elements are further classified as

1) Inert gases
2) Representative elements
3) Transition elements
4) Inner transition elements.

INERT GASES :

- Inert gas elements have all completed shells.
- They belong to '0' group in the periodic table.
- Helium, Neon, Argon, Krypton, Xenon, Radon are called noble gases or rare gases or inert gases (or) aerogens

- These are monoatomic gases.
- They are chemically inactive.
- They are placed at the extreme right of the periodic table.

REPRESENTATIVE ELEMENTS:

- s-block or p-block elements except '0' group are called representative elements.
- They have only one incomplete outer shell.
- These elements attain the nearest inert gas configuration by losing or gaining or sharing electrons.
- They are chemically active.
- A few metals and metalloids are found in representative elements.
- Because of their reactivity and frequent occurrence they are called representative elements.
- They include most reactive metals and most reactive non-metals.

TRANSITION ELEMENTS:

- These are d-block elements.
- They have two incomplete outer shells ultimate and penultimate.
- Their general electronic configuration is $(n-1)d^{1-10}\ ns^{1-2}$.
- Neutral atoms or Ions having incomplete
- d-orbitals are called transition elements.
- Zn, Cd and Hg are not considered to be transition elements as their atoms and Ions have completed d-orbitals.
- Small atomic size, high nuclear charge, and unpaired d-electrons give characteristic properties to transition elements.
- Transition elements are hard and dense metals.
- They have high melting and boiling points.
- They are good conductors of heat and electricity.
- They show variable oxidation states.
- They form coloured compounds.
- They form complexes or co-ordinate covalent compounds.
- They readily form alloys like brass, bronze, german silver etc.

INNER TRANSITION ELEMENTS:

- The f-block elements are called inner transition elements as they bring about transition among transition metals.
- The differentiating electron enters into the
- f-orbital of anti penultimate shell.
- These elements have three incomplete outer shells.
- The general electronic configuration of these elements is $(n-2)f^{1-14}\ (n-1)d^{0-1}ns^2$.
- These elements show similar properties due to the similar electronic configuration in the last two shells.
- They exhibit the common oxidation state + 3.
- All the inner transition elements belong to the same group (i.e. IIIB)

PERIODICITY-PERIODIC PROPERTIES:

- Repetition of properties after intervals of atomic number values 2, 8, 18 and 32 is called periodicity.

The properties are called periodic properties.

ATOMIC RADIUS:

- The distance between the centre of the nucleus and electron cloud of the outermost energy level is called the atomic radius.
- It cannot be measured directly.
- It can be measured from the inter nuclear distance of bonded atoms using x-ray diffraction techniques.
- The atomic radius depends on factors like
  1) the number of bonds formed by atom
  2) nature of bonding
  3) oxidation state

CRYSTAL RADIUS OR ATOMIC RADIUS:

- It is applicable for metal atoms.
- Half of the distance between the nuclei of two adjacent atoms in metallic crystal is called crystal radius.
- It is measured in angstrom units
  (1 Å = 10^{-8} cm = 10^{-10} m) or nanometres.
  (1 nm = 10^{-9} mm = 10^{-7} cm)

COVALENT RADIUS:

- This is used for non metals which form covalent bonds.
- It is half of the distance between the nuclei of two atoms connected by a covalent bond in a homodiatomic molecule.
- Crystal radius is slightly greater than the covalent radius.

VANDERWAAL'S RADIUS:

- It is used for molecular substances in the solid state only.
- It is half of the distance between the nuclei of two adjacent nonbonded atoms in neighbouring molecules.
- Vanderwaal's radius is greater than the covalent radius as the Vanderwaal's forces are weak.
- Vanderwaal's radius is approximately 40% greater than the covalent radius.

TREND :

- The atomic radius increases down the group due to the addition of new shells and increase in screening effect.
- The atomic radius decreases from left to right in a period due to the increase in nuclear positive charge.
- In a given period alkali metal is the largest atom and halogen is the smallest atom.
- Since the inert gas atoms are non bonded their atomic radius should be taken as the Vanderwaal's radius.
- Thus in every period inert gas atom is larger than halogen.
- In 2^{nd} period inert gas atom is the largest and in other periods alkali metal atom is the largest.
- The decrease in atomic radius of transition elements is less than expected due to the screening effect of (n – 1)d electrons.

- Thus transition metals have similar atomic radii.
- In transition families, atomic radius increases normally from 3d to 4d series. But there after it doesn't charge much due to lanthanide contraction.

LANTHANIDE CONTRACTION:

- In inner transition elements the differentiating electron enters into 'f' orbitals of the antepenultimate shell.
- As the atomic number increases in lanthanides due to the dispersed shape of f-orbitals and their poor shielding effect the atomic and ionic radii steadily decrease. This is called lanthanide contraction.
- Lanthanide contraction is also observed in 5d transition series.
- The atomic radius of 5d transition elements are very close to those of 4d transition elements due to Lanthanide contraction.
- As a result 4d and 5d transition elements are more similar in properties when compared to 3d and 4d transition elements e.g. Zr and Hf resemble most closed to each other than other elements.

IONIC RADIUS:

- The distance between the nucleus and the outermost e^- in an ion.
- When an atom loses one or more electrons a positive Ion is formed.
- The cation is smaller in size then the neutral atom.
- The ionic radius is smaller than the atomic radius.
- As the number of electrons removed from the atom increases, the ionic radius is further decreased.

	ATOMIC RADIUS	**IONIC RADIUS**
i) Na	1.86 Å	Na^+ 1.02 Å
ii) Fe	1.17Å	Fe^{2+} 0.76 Å e^{3+}0.64 Å

- If an atom gains electrons negative ion is formed (anion).
- The negative ion is bigger in size than the neutral atom.
- The size of anion increases with increase in negative charge.
- With increase in z/e ratio ionic radius decreases.
- In cations, the z/e ratio is greater than 1 and in anions, the z/e ratio is less than 1. (charge per electron = z/e)
- The atomic radius of chlorine atom (0.99 Å) is much smaller than that of chloride ion (1.81 Å).

ISOELECTRONIC SPECIES:

- Ionic species having the same number of electrons is called isoelectronic species.
- In isoelectronic series the size decreases with the increases in nuclear positive charge.

Ion	C^{4-}	N^{3-}	O^{2-}	F^-	Na^+	Mg^{2+}	Al^{3+}	Si^{4+}
Ion Radius (Å)	2.60	1.71	1.40	1.36	0.95	0.65	0.50	0.41

- In isoelectronic series, size decreases with increase in number of protons.

i) Smallest atom is 'H' ii) Largest atom is Fr

iii) Smallest cation is H^+ iv) Smallest anion is H^-

v) Largest cation is Cs^+ vi) Largest anion is I^-

IONIZATION POTENTIAL OR IONIZATION ENERGY:

- The minimum amount of energy required to remove the most loosely bound electron from an isolated gaseous atom is called its first ionization potential (I_1). $M(g) \xrightarrow{I_1} M^+_{(g)}$
- The energy required to remove an electron from a unipositive ion is called its second ionization potential (I_2). $M^+_{(g)} \xrightarrow{I_2} M^{+2}{}_{(g)}$
- The number of ionisation potentials of an atom is equal to it's atomic number.
- The second ionization potential is greater than the first ionization potential.
- With the successive removal of electrons ionization potential increases due to increased nuclear charge.
- $I_1 = 13.6 \text{ X } z^2$ (z = effective nuclear charge)
- Ionization potential is measured in eV/atom and ionization energy is measured in kJ/mole (1 eV/atom = 96.45 kJ $mole^{-1}$ or 1 eV/atom = 23.06 kcal/mole)
- Ionization energies are determined from spectral studies as well as from discharge tube experiments.
- Ionization potential depends on the following factors.
  1) Atomic radius.
  2) Nuclear positive charge
  3) Screening effect or shielding effect.
  4) Extent of penetration of valence electrons.
  5) Completely or half-filled sub shells.
- lionization potential decreases as the atomic radius increases.
- Ionization potential increases as the nuclear positive charge increases.

SCREENING EFFECT:

- The protection given by the inner electrons to the valence electrons from the nucleus is called screening effect.
- As the inner electrons increase in number, screening effect increases and the ionization potential decreases.
- The extent of penetration of orbitals towards the nucleus follows the order s > p > d > f for a given principle quantum number n.
- With increase in the extent of penetration, I.P. increases. s > p > d > f.
- Atoms having completely filled and half filled orbtials are more stable and need more energy for ionization.
- Be has greater ionisation potential than B due to the completely filled s-orbital.
- Nitrogen has greater ionization potential than oxygen due to the half filled p-orbitals.

  I_1 of 2^{nd} period: Li < B < Be < C < 0 < N < F

  I_2 of 2^{nd} period: $Be^+ < C^+ < B^+ < N^+ < F^+ < 0^+ < Li^+$

TREND OF IONIZATION POTENTIAL:

- Ionization potential decreases down the group due to increase in size and increases in screening effect.
- lonisation potential increase in a period form left to right due to decrease in size and increase in nuclear charge.
- In any given period, alkali metal has less. I.P. value and inert gas has highest I.P. value.

- In transition series I.P. value slowly ↑ due to less screening effect of (n – 1)d electrons.
- Alkali metals have low I.P values. Lowest I.P. value is for Cs.
- Noble gases have high I.P values. Highest I.P. value is for He.
- In transition families I.P. value ↓ from 3d to 4d and slightly increases there after due to lanthanide contraction.
- The I_1 values of 1st few elements:
- He > Ne > F > Ar > N > Kr > 0 etc.
- I.P. curve is obtained by plotting I.P. Vs atomic number.
- In I.P. curve, peaks are occupied by Inert gases.
- Troughs are occupied by alkali metals.
- Ascending portions are occupied by II-A group metals.
- Descending portions are occupied by VII-A group elements.

ELECTRON AFFINITY:

- The amount of energy released when an electron is added to a neutral gaseous atom is called electron affinity.
- Energy is released when the first electron is added to the neutral atom.
- The energy is required to add the second electron or e^- s *to* uninegative ion.
- Therefore for most of elements E_1 values are negative (energy released) and for all the elements E_2, E_3 etc values are positive (energy absorbed).
- Even E_1 values are positive for some elements due to their stable configurations. E.g.: 'O' group elements, Be, Mg, N.
- Numerically I_1 of atom 'M' is equal to E_1 of M^+ ion.
- Numerically E_1 of atom 'M' is equal to I_1 of M^- ion.
- They cannot be determined directly.
- They are calculated indirectly using Born-Haber cycle.
- Electron affinity is measured in kJ mol^{-1} or eV / atom.
- Electron affinity depends on
  (i) **Size:** E.A increases with decrease in size
  (ii) **Nuclear charge**: E.A. increases with increase in nuclear charge
  (iii) **Screening effect**: EA decreases with increase in screening effect.
  (iv) **Electronic configuration**: Atoms with stable electronic configuration have zero or negative electron affinities.
  (v) **Number of valence electrons in valence shell.** E.A increases with increase in the number of valence e^-s.

TREND:

- E.A. decreases from top to bottom in a group due to increase in size and increase in nuclear charge.
- But in most of the groups, the 1st element has abnormally low E.A. value than that of the remaining elements, eg. F < CI, O < S.
- EA ↑ from left to right in a period due to ↓ in atomic size and ↑ in nuclear charge.
- In every period halogen has highest E.A value.
- E_1 trend of few elements: CI > F>Br>I>S>Si.

APPLICATIONS OF E.A.:

(i) We can predict the ability of atom into anion.

(ii)Oxidising power of element.

(iii) We can predict the tightness of binding an electron to the valence shell.

ELECTRONEGATIVITY:

- The tendency of an atom to attract the shared electron pair towards itself in a molecule is called its electro negativity.
- Electron affinity is the property of an isolated atom and electronegativity is the property of a bonded atom.
- Electron affinity is absolute phenomenon electro negativity is relative phenomenon.
- Electron affinity has units and electronegativity has no units.
- Electron affinity of an element is fixed and electronegativity is slightly variable.
- Electronegativity depends on
  i) **Size of atom:** EN. decreases with increase in size of atom
  ii) **Nuclear charge***:* E.N. increases with increase in nuclear charge
  iii) **Screening effect:** E.N. decreases with increase in screening effect.
  iv) **Number of valence electrons:** E.N increases with increase in number of valence e^-s.
  v) **S-character of hybrid orbital:** E.N. increases with increase in S-character of hybrid orbital.

TREND:

- Electronegativity decreases with increase in size and increase in screening effect in a group from top to bottom.
- Among all groups, I-A group has least E.N. and VII-A group has highest E.N. Values.
- Electronegativity increases with decrease in size and increase in nuclear charge in a period from left to right.
- In a given period I-A element has least E.N. and VII-A element has highest E.N. value.
- 'O' group elements have zero E.N values.
- E.N of the 1^{st} few elements:
  F > O > N ≈ CI > Br > C ≈ I > H
- In Pauling's scale electroegativities are calculated from the bond energies.

SCALES OF ELECTRONEGATIVITY :

Pauling scales :

- It is widely accepted scale.
- $X_A - X_B = 0.208.\sqrt{\Delta}$ if Δ is in kcal/mole or
- $X_A - X_B = 0.1017\sqrt{\Delta}$ if Δ is in kJ/mole.
- X_A and X_B are the electronegativities of two elements A and B.
- Δ-difference between calculated and experimental bond energies of A – B bond.
  Δ = ionic - resonance energy
  Δ = polarity of A - B bond.
- In the determination of electronegativities of other elements. Hydrogen is reference element.
- Fluorine is the most electronegative and Cs is the least electronegative.
- Electronegative value of Hydrogen is 2.1.

MULLIKAN SCALE:

- Mullikan suggested that the electronegativity of an element is the average of its ionization energy and electron affinity.

$$E.N = \frac{\text{ionization energy} + \text{electron affinity}}{2} \text{ (eV/atom)}$$

(E.N = Electronegativity)

- These values are about 2.8 times greater than the values on Pauling's scale.
- The commonly accepted electronegataivity values are obtained from the formula.

$$E.N = \frac{\text{ionization potential} + \text{electron affinity}}{544} \text{ (kJ/mole)}$$

- Mulliken scale is not much useful due to the following drawbacks.

i) It is applicable only for monovalent elements.

ii) As per Mulliken scale, Inert gases should posses higher E.N values.

Conversion:

$$1 \text{ Unit of pauling} = \frac{1 \text{ unit of Mulliken}}{2.8}$$

APPLICATIONS OF ELECTRONEGATIVITY:

i) If the electronegativity difference of the bonded atoms is greater than or equal to 1.7, the bond is ionic. If the electronegativity difference is less than 1.7 the bond is covalent or polar covalent.

ii) Electronegativity values are useful in writing the formulae of compounds.

iii) In calculation of oxidation states.

iv) In predicting oxidising power of elements.

VALENCY :

- Valency was introduced by mendeleef.
- Valency is the combining capacity i.e. number of bonds formed by an atom.
- Valency of an element can be defined as the
  - i) number of hydrogens with one atom of element combines.
  - ii) number of chlorine atoms with which one atom of element combines.
  - iii)double the number of oxygen atoms with which one atom of element combines.
- Valency of a given group element will be same in general.
- Highest valency of an element will never exceed it's group number.
- Valency of an element may be equal to it's group number or 8 – group number.
- Valency of a given period elements will be different and it changes by 1 unit from one element to next element.
- w.r.t. oxygen. Valency increases from 1 to 7 in a period.
- w.r.t. hydrogen, valency increases from 1 to 4 and then decreases towards the end of the period.
- For zero group elements valency is zero
- Valency is a whole number without positive or negative sign.

OXIDATION STATES:

- The charge which an atom appears to be possessed in a molecule or ion is called its oxidation state.
- Oxidation state may be positive or negative or a fraction or zero.

- For s-block elements oxidation number is equal to the group number.
- p-block elements show different oxidation states which often differ by 2 units.
- The s-electron pair in the valence shell of the heavier elements of p-block show reluctance in bond formation. This is called inert pair effect.
- In p-block group III elements have a common oxidation state of +3.
- For thallium + 1 oxidation state is more stable than + 3 oxidation state, due to inert pair effect.
- Group IV elements show + 4 and +2 oxidation states.
- Group V elements show + 5 and +3 oxidation states.
- The common oxidation state shown by group VI elements is –2.
- The other oxidation states shown by VI group elements are + 2, +4 and +6.
- Halogens are the most electronegative elements and show the common oxidation state –1.
- Fluorine always shows the oxidation state–1.
- The other halogens show positive oxidation state +1, +3, +5 and +7.
- The common oxidation state of transition or d-block elements is +2 due to the ns^2 electronic configuration.
- Ruthenium (Ru) and osmium (Os) show the highest oxidation state of +8.
- Carbon show the lowest oxidation state of–4.
- f-block elements or inner transition elements show the common oxidation state of +3.
- The highest oxidation state shown by any element is equal to its group number.

ELECTROPOSITIVE NATURE:

- The tendency of an element to lose an electron is called its electropositivity.
- Electropositivity is a metallic character.
- Smaller the ionization potential greater the electropositive character.
- More electropositive metals form ionic compounds.
- Strongly electropositive metals readily liberate hydrogen from water and dilute acids.
- More electropositive metals form strong basic oxides and hydroxides.
- The ions of strong electropositive metals do not undergo hydrolysis.
- Electropositive nature increases down the group and decreases along a period.
- Alkalimetals are the most electropositive metals.
- Strong electropositive elements are strong reducing agents.

Metallic and non metallic nature:

- Metallic nature means electropositive nature
- Most electropositive or most metallic is Cs.
- Non-metallic nature means electronegative nature.
- Most electronegative or most non-metallic is F.
- Metallic nature increases and non-metallic nature decreases down the group.
- Metallic nature decreases and non-metallic nature increases along a period.

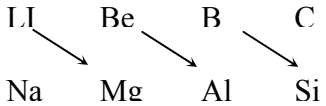
Acidic and basic nature of oxides:

- In general, metal oxides are basic and dissolve in water to form hydroxides. e.g.: Na_2O, K_2O, CaO, BaO.
- Alkali metal oxides are the most basic oxides.

- Oxides of nonmetals are acidic and dissolve in water to form acidic solutions e.g.: CO_2, SO_2, P_2O_3, Cl_2O_7 etc.
- Oxides of halogens are the most acidic oxides.
- Oxides which show both acidic and basic properties are called amphoteric oxides.
- Metalloids and a few metals form amphoteric oxides. e.g.: BeO, ZnO, Al_2O_3, SnO, SnO_2, PbO, PbO_2, As_2O_3, Sb_2O_3.
- Basic nature of oxides increases and acidic nature decreases down the group.
- Across a period acidic nature increases and basic nature decreases in oxides.
- The 2nd period element which forms most acidic oxide is nitrogen (N_2O_5)
- The 3rd period element which forms most acidic oxide is chlorine (Cl_2O_7)
- Among all, the most acidic oxide is CI_2O_7 and most basic oxide is Cs_2O.
- Fluorine does not form oxides, instead it forms fluorides eg. O_2F_2 and OF_2.

DIAGONAL RELATIONSHIP:

- The lighter elements of short periods show similarities with the diagonally arranged ones. This is called diagonal relationship.



- This is found between 2nd and 3rd period elements.
- Diagonal relation is weaker than group similarities.
- Diagonal relation disappears beyond IV-A group.
- Diagonal relation is due to
  i) similar ionic sizes and similar electronegativity values.
  ii) similar polarising power.

$$\text{Polarising power} = \frac{\text{Charge}}{(\text{radius})^2}$$

Trend of Melting and Boiling Points:

- In a period MP's and BP's will increases first and decreases towards the end of period.
- In a group MP's and BP's will decreases from top to bottom in IA, IIA, IIIA, IVA.
- In VA, VIA, VIIA and zero groups, MP's and BP's will increases from top to bottom.

TREND OF ATOMIC VOLUME :

- Atomic volume is the volume occupied by 1 gram – atom of element.
- In a group atomic volume increases from top to bottom.
- In a period atomic volume first decreases and then increases towards the end.

TREND OF DENSITY:

- In a group density increases from top to bottom due to increase in atomic mass.
- In a period density increases from left to right due to decrease in atomic size