

Periodic Properties

Periodicity is related to the periodic properties associated with the elements in a periodic table. Periodicity occurs due to the repetition of similar outer electronic configuration of different elements. There will be a variation in the properties of the elements which will repeat after certain intervals of time (either they will increase / decrease).

Periodic Table is defined as the arrangement of elements according to their atomic number. It is also known as modern Periodic Table. It is given by Moseley.

Periodic Table is divided into 4 blocks, 7 periods and 18 groups.

| | |
|---------|--------------------------------------|
| s-block | → Group 1 & Group 2 (Representative) |
| p-block | → Groups 13 to 18 |
| d-block | → Groups 3 to 12 (Transition) |
| f-block | → Lanthanoids & Actinoids |

Penetration of Orbitals:-

Penetration power of an orbital is defined as the ability to release energy to accept electron. It depends on the closeness of the orbital towards the nucleus.

More the closeness towards the nucleus more will be the release of energy. It is so because, the inner shell electron shield nucleus more than the outermost shell i.e.

Pertacting power of one in subshells

$$s > p > d > f$$

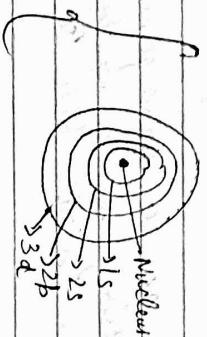
Pertacting power of an electron

$$1s > 2s > 2p > 3s > 3p > 4s > 3d > 4p > 4f$$

and energy of an e⁻ for each shell and subshell

$$1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p$$

Shell model of atom



The e⁻ are -vely charged and are pulled towards the truly charged nucleus. The e⁻ are thus attracted to the nucleus, but at same time they repel each other.

The repulsion which one -field \rightarrow Shielding
The attraction which an e⁻ feels \rightarrow Pertraction

Electronic Configuration of an Element:-

It is the representation of electronic arrangement in different shells or subshell of an atom.

Rules that governs electronic configuration.

Afflans Principle: afflans' Greek meaning is 'Building up'

Or 'to setup'. This rule states that filling of e⁻ in as orbitals occurs from lowest energy to

Highest Energy shell:

$$\begin{array}{c} 1s \\ 2s \quad 2p \\ 3s \quad 3p \quad 3d \\ 4s \quad 4p \quad 4d \quad 4f \\ 5s \quad 5p \quad 5d \quad 5f \end{array}$$

Quantum Numbers: It gives complete information of an electron.

These are of four types:-

- (a) Principal quantum number (n) \rightarrow shell
- (b) Azimutal quantum number (l) \rightarrow sub-shell
- (c) Magnetic quantum number (m) \rightarrow + l to - l
- (d) Spin quantum number (s) \rightarrow 1/2 (+1/2 or -1/2)

for s-orbital \rightarrow l=0

p \rightarrow l=1

d \rightarrow l=2

f \rightarrow l=3

- Pauli's Exclusion Principle: It states that no two e⁻ in an atom can have the same set of four quantum nos.

- Mendel's Rule of Multiplicity: It states that if filling in an orbital occurs in such a way that every subshell must be slightly filled first only then next e⁻ enters in second subshell.

$(n-1)$ \rightarrow Multinucleate shell
 $(n-2)$ \rightarrow Antiperimultinuclear shell

Classification of Elements into s,p,d,f Blocks:

The classification depends upon the type of orbital into which the last e- of atom enters.

a) S-block elements - In this, last e- enters the s-orbital of their outermost energy level.

It consists of group 1 and group 2

e.i.c of s-block elements = ns^{1-2}

$ns^1 \rightarrow$ alkali metals

$ns^2 \rightarrow$ alkaline earth metals

b) p-block elements, similarly, last e- enters in p-orbitals.

It consists of elements of group 13-18.

General e.i.c = $ns^2 np^{1-6}$

c) d-block elements \rightarrow last e- enters in d-orbital.

It consists of elements of group 3-12 and are also known as transition elements.

General e.i.c = $(n-1)d^{1-10} ns^{1-2}$

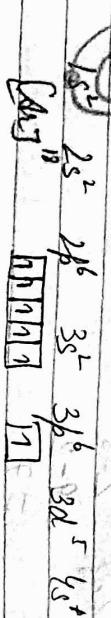
$Cr \rightarrow 24$ [18] 1s² 2s² 2p⁶ 3s² 3p⁶ 3d⁵ 4s²

[18] [12] [12] [12] [12] [12]

d) f-block elements \rightarrow last e- enters in f-orbital of their atoms. It consists of two series lanthanides & actinides and are known as inner transition elements.

General e.i.c = $(n-2)f^{1-4} (n-1)d^{0-2} ns^2$

So, stable e.i.c of Cr is



| | | |
|-----------------|---|----------------|
| Shade | d-Block $(n-1)d^{10} ns^2$ | p-block ns^1 |
| ns ² | d-Block elements $(n-2)f^{1-4} (n-1)d^{0-2} ns^2$ | |

Division of periodicic table into s,p,d,f elements

Example: a) e.i.c of F E-

F, Z = 9 1s² 2s² 2p⁵

[18] [12] [12] [12] [12]

b) Electronic configuration of Se

Se, Z = 10

[18] [12] [12] [12] [12]

This e.i.c is not valid when an element wants to have stability we know that half filled or fully filled e.i.c are stable in nature.

Now for Cr to have stability, it is required only 1e- in 3d, still to become half-filled stable state, which will be replaced from 4s-shell.

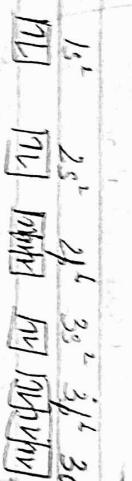
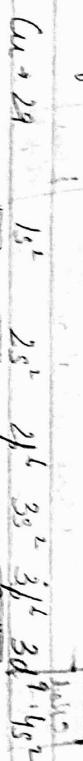
So, stable e.i.c of Cr is



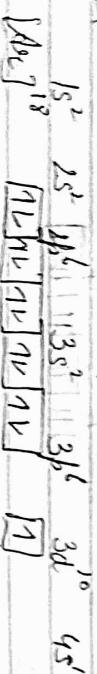
Effective nuclear charge \rightarrow Actual charge left by valence e⁻
shielding / screening effect - Repulsive force left by valence e⁻ from nucleus.

\rightarrow Variation of shielding of orbital energies.

c) Z_{eff} of Cu



This Z_{eff} is not valid when an element wants to have stability. We know that fully filled Z_{eff} is stable in nature. Now for Cu to have stability it requires only 1e⁻ in 3d shell to become fully filled. That's why Z_{eff} will be replacing from 3d shell.

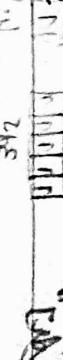


Effective Nuclear Charge: It is the net effective charge experienced by the e⁻ after the shielding / screening effect of inner shells. It is denoted as Z_{eff} and given by

$$Z_{\text{eff}} = Z - S$$

where Z = atomic no / total nuclear charge
S = screening constant / shielding effect of inner shells.

Example:- Z_{eff} of F



i) Valency: It is determined by the no. of e⁻ in the outermost shell. And outermost shell e⁻ are exposed to all valence shell & so highly valence e⁻.

Variation of Valency in periodic table:-

| | |
|-----------|-----------------|
| For e.g., | Li Be B C N O F |
| Valency | 1 2 3 4 3 2 1 |

for e.g., all elements of group 1 have valency 1 and group 2 have same valency 2.

2) Atomic and Ionic sizes:-

Atomic radius is defined as the distance from the centre of the nucleus to the outermost shell containing. It represents atomic size.



Types of atomic radius

Covalent radius

Van der Waals radius

Ionic radius

- a) Covalent Radius: It is defined as one-half of the distance b/w the nuclei of two covalently bonded atoms of same element.

Equivalent = $\frac{1}{2}$ [Bond length]

i) r₀



constant radius

ii) r₁



Variation of Ionic Radii:-

- Along Period:- Ionic radii (covalent and Van der Waals) decrease with increase in atomic no. as we move from left to right.

Atomic nuclear charge increases because new e⁻ are entering. The new e⁻ enters into same subshell which increases nuclear charge & hence attraction of e⁻ towards nucleus gets increase.

Nuclear charge \propto

Attraction \propto

Atomic size \propto

Variation of Ionic Radii:-

- b) Vanderwaals Radius: Vanderwaals forces are the weak forces which exists among two non-bonded molecules.
- for ex:- Intermolecular force in Br_2 & AlCl_3 molecule.

$\text{Cl} - \text{---} \text{Cl}$, $\text{Na} - \text{---} \text{Cl}$, $\text{Na} - \text{---} \text{Na}$

Now, molecular structure

(i) r₀



(ii) r₁



It is defined as the distance from the center

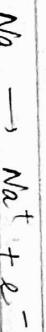
- of the nuclei of two non-bonded isolated atoms, belonging to two neighbouring molecules of an element in the solid state.

- c) Ionic Radii: Ionic radii is present in those molecules which are having ionic bond in them for eg., formation of ionic bond is in which Na⁺ cations bind to anions.

Ionic radii is defined as "in term of anion and cation, distance from the center of cation and anion to the center at which both have exert's their effect."

- Variation along group - As we move down a group, the ionic radius increases with increase in atomic number and principal quantum level also increases.
- Comparison b/w size of cation and anion from its parent atom.

Na and Na^+



\rightarrow size of cation (Na^+) is always less than its parent atom because the nuclear charge increases and oxidation state decreases from 0 to +1 in terms of e^- .

Due to increase of nuclear charge, attraction towards the nucleus also increases and hence size get decreases.

\rightarrow Size of anion is more than that of its parent atom because due to e^-e^- repulsion the nuclear charge will decrease but due to shielding effect of inner shell & the distance of outermost e^- from center of nucleus will get increases \therefore size increases.



3) Ionisation energy - It is defined as the amount of energy required to remove the most loosely bound e^- from an isolated gaseous atom



It gives a measure of the ease with which an electron can lose an e^- and change into cation. It is expressed in terms of kcal/mol (OR) kJ/mol (OR). It is defined as initial energy required to remove an e^- from neutral gaseous atom. Amount of energy depends on which factors on which I.E. depends. Factors formed from neutral atom

i) Size of atom: - I.E. depends upon distance of e^- and nucleus.

As size of atom increases, attraction e^- are less strongly attracted by the nucleus \therefore force of attraction is inversely proportional to square of the distance of e^- charged particle. As a result, I.E. decreases with increase in size of atom.

ii) Charge on nucleus: - The attractive forces b/w the nucleus and the e^- increases with increase in nuclear charge. \therefore force of attraction is directly proportional to product of charges on nucleus \therefore with increase in nuclear charge, it becomes more difficult to remove an e^- and hence I.E. increases.

iii) Screening effect of inner e^- : - If no of e^-e^- pair is less, then I.E. will decrease.

4) Penetration effect of e^- : - $S-e^-$ are more penetrating towards the nucleus than $p-e^-$, so the order will be

$$S > p > d > f$$

So if penetration of e^- increases, it will be short

nucleus, so T.E will be high with means T.E increase with increase in penetrating power of it.

Variation of T.E:

- Along Period:- T.E increases with increasing atomic no. in a period. So on moving across a period from left to right, nuclear charge increases and there is decrease in atomic size, the electron is more tightly held by the nucleus. i.e. more e⁻ is needed to remove it and hence T.E decreases. increase

- Along a Group:- Within a group, there is a gradual decrease in T.E in moving from top to bottom. So going from top to bottom, nuclear charge increases, atomic size also increases and so shielding effect on outermost e⁻ also increases due to increase in no. of inner e⁻.

- Electronegativity:- It is defined as the tendency of an atom to attract shared paired e⁻ towards itself.

Variation of Electronegativity:-

- Along a Period:- On moving along a period the electronegativity increases due to the decrease in atomic size and nuclear charge increases.
- Effective nuclear charge increases.

| for eg - | Li | Be | B | C | N | O | P |
|--------------|----|----|---|---|---|---|---|
| EN increases | | | | | | | |

• Along a group:- Atomic size increases when we move from top to bottom and due to this repulsive increases, so electronegativity decreases.

| for eg:- | F | Cl | Br | I |
|----------|---|----|----|---|
| EN | | | | |

factors affecting the electronegativity:-

1) Oxidation state:- O.S & EN

Electronegativity increases as the oxidation state of atom increases. (tendency to attract the e⁻ will increase for Anions, EN decreases with increasing -ve charge of ion).

2) State of hybridization, this depends upon penetrating power:

S-orbitals are nearer to nucleus than p, d, f orbitals because of this higher penetration power is s-orbitals will have greater e⁻ attracting power of electronegativity.

(Or)
EN increases with increasing s-character of hybrid orbital.

3) Nature of substituents:- Electronegativity of group varies with the nature of substituents. It is due to the Inductive effect of substituent groups.

- (i) Size of atom :- Atoms with small size have higher values of electronegativity. This is so because, smaller atoms have higher effective nuclear charge (Z_{eff})
- (ii) Shared pair of e⁻ is pulled more strongly by the nucleus and electronegativity is high.

Applications of Electronegativity:-

- determination of oxidation state
- state of hybridization
- calculation of Bond Order
- determination of dipole moment

Polarizability :-

Note :- A cation approaches an anion, the electronic cloud of anion will get distorted due to the polarisation by cation.

The power of cation to distort an anion is known as polarising power / polarizability. The tendency of anion to get distorted is known as polarization.

For ex:- (i) Na + Cl \rightarrow Na⁺ Cl⁻ (Ionic Bond)

(ii) O^2- Before Polarization

(iii) O^2- Polarization occurs hence covalent behaviour occurs.

Mathematically,

$$\chi = \frac{\mu}{kE}$$

where χ is polarisability constant (Cm^2V^{-1})
 μ is induced dipole
 E is strength of electric field

Fajan Rule:- This rule shows the covalent behaviour of molecule and covalent behaviour depends upon polarization.

(a) Size of Cation \Rightarrow

Size of Cation \propto Polarization / Covalent Behaviour

Smaller the cation greater will be the polarisability and larger is the covalent behaviour.

Li+ Na+ K+ Rb+ Cs+

Covalent behaviour \downarrow

Polarization \uparrow

(b) Size of Anion:-

Li- Na- K- Rb- Cs-

Size of Anion \propto Covalent Behaviour

Covalent behaviour \uparrow

(c) Electronic Configuration:- Fully filled and half filled electron configuration have no effect on polarisability because of s.p. hybridization

for noble gas, configuration of cation produces better shielding and less polarizing power.

b) Electron affinity :- The amount of energy released when an e^- is added to neutral isolated gaseous atom.



When one e^- is removed from an atom energy is required for the process of removal of e^- but the energy is released when electron is added to neutral atom. Electron affinity is expressed as kJ mol^{-1} .

Factors affecting Electron Affinity (EA) Electron gain Energy:

Q. EA of fluorine is unexpectedly less than that of chlorine.
Ans. low value of E_A of atom is due to very small size of atom and due to small size, they have strong inter-electrostatic pull and thus incoming e^- does not feel much attraction.

On other hand, electron affinity of Cl atom is larger than F atom where e^- is added to large 3p orbital, which can easily accommodate additional e^- .

Halogens have highest E_A

Ans. This is so because halogen have general e.c. $n s^2 n p^5$ and have only one e^- less than stable noble gas ($n s^2 n p^6$) configuration. Thus they have very strong tendency to accept additional e^- and so E_A are high.

- Nuclear charge :- Nuclear charge $\propto EA$. As nuclear charge increases, E_A also increases. And also attraction to incoming e^- increases.
- Electron Configuration :- Elements which have stable e.c. of half filled valence subshell have very small tendency to accept e^- and so E_A are low or almost zero.

Variation of Electron Affinity

Along a group:- On moving from top to bottom, distance along and nuclear charge decreases. So E_A decreases.

Along Period:- On moving from left to right, atomic size decreases and nuclear charge increases. So E_A increases.

7.) Oxidation State / Oxidation Number -

It is defined as the charge exhibited by an ion when it is formed from its neutral atom by loss of one or more e^- from atom.



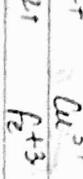
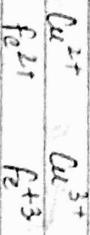
In s-block, this gives no. of their oxidation number.

In p-block, oxidation state is either given no. of 8 - total no. of valence e⁻.

$$\text{for } \text{Cl}^- \quad F, Z=9(1,7)=\cancel{8}-1 \quad 7-8=-1$$

$$\text{SO}_3, Z=16(2,6,3)=$$

d-block elements, having variable oxidation state



f-block elements, show very high and variable osr
 La³⁺, La⁵⁺, La⁷⁺, La⁹⁺

Example:- Calculate O.S. of A in K_2CuO_2

$$2+2\times 2-14 = +6$$

Fe sulfate [Co₃]
 DS = 0

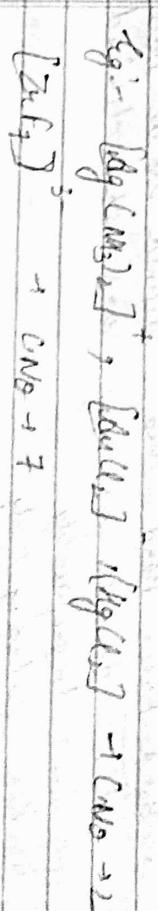
O.S. of Acids and Bases

| | | |
|--------------------------|-----------------|-------------------|
| $\text{Co} = 0$ | $\text{H} = +1$ | $\text{Cl} = -1$ |
| $\text{H}_2\text{O} = 0$ | $\text{O} = -2$ | $\text{SCN} = -1$ |
| $\text{NH}_3 = 0$ | $\text{N} = -3$ | |

Co-ordination Number - It is defined as the no. of atoms, ions or molecules that central atom holds in its nearest in a complex.

(OR)

It is defined as the total no. of univalent ligand in a coordination sphere when metal and ligands form coordinate bond.



Hard and Soft acids and Bases (HSAB)

It is given by Ralf Pears to explain the stability of complexes. It is also referred to as Pearson concept.

According to this principle Hard acids prefer to combine with hard base and soft acids prefer to combine with soft base.

Now according to Pearson theory of hard acids, hard

bases and soft acids, soft base is explained by Lewis acid and Lewis base.

According to Lewis concept an acid is a substance which accept e⁻ and a base is a substance which donates e⁻.

Hard Acid:- It is defined as acid which accept e⁻ they are characterised by:-

- (i) High +ve charge
- (ii) Small size
- (iii) Soft easily polarisable
- (iv) Strongly solvated
- (v) Empty orbitals in valence shell
- (vi) With high energy LUMO



Hard Base:- Hard base is defined as the base which donates e⁻. They are characterised by:-

- (i) Low Polarizability
- (ii) High Electronegativity
- (iii) Not easily oxidised
- (iv) Completely filled Atomic Orbitals
- (v) With low energy LUMO



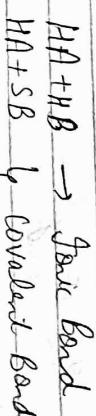
Soft Acids:- Soft acids is defined as the acid which don't accept e⁻. They are characterised by:-

- i) Low +ve charge
- ii) Large size
- iii) Highly polarisable
- iv) Completely filled A.O with large size
- v) With low energy LUMO



Soft base:- Soft base is defined as the base which don't accept e⁻. They are characterized by:-

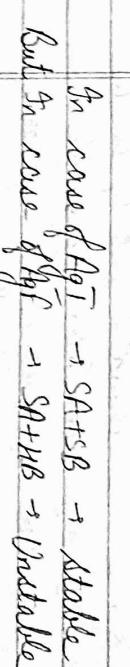
- (i) Small size
- (ii) High polarisability
- (iii) Low electronegativity
- (iv) Highly oxidised



Q Out of Silver iodide (AgI) and silver fluoride which is more stable and why?

AgI is more stable than AgF :: Stability depends upon NSA principle.

According to which Hard acids prefers to combine with hard base and soft acids prefers to combine with soft base.



But In case of $\text{AgF}^- \rightarrow \text{SA} + \text{HB} \rightarrow \text{Unstable}$

Stability also depends upon size of cation. more is charge of cation, lesser is the ionic radius so more is the attraction, higher will be the stability of complex.