CLASSIFICATION OF ELEMENTS AND PERIODICITY IN PROPERTIES

PERIODIC TRENDS IN PROPERTIES OF ELEMENTS

ATOMIC PROPERTIES OR PERIODIC PROPERTIES OF ELEMENTS

The physical properties of the elements can be divided into two categories :

(i) Properties of individual atoms: Properties such as atomic and ionic radiate ionisation enthalpy, electron gain enthalpy, valency, screening effect, effective nuclear charge and electronegativity are the properties of individual atoms and directly dependent on their electronic configurations.

(ii) Properties of group of atoms: Properties such as melting point, boiling point, heat of fusion, heat of vaporization, density, atomic volume, etc., are the properties of group of atoms together and are indirectly related to their electronic configurations. All these properties which are directly or indirectly dependent on the electronic configuration of the elements are called atomic properties.

Since electronic configurations of the elements are periodic function of atomic numbers of the elements, therefore, these atomic properties are also a periodic function of atomic numbers of the elements. On account of this, atomic properties are also called periodic properties.

Valency: It is defined as the combining capacity of the elements. The word valency is derived from an Italian word "Valentia" which means combining capacity.

Old concept: Given by: Frankland

Valency with respect to Hydrogen: Valency of H = 1 It is defined as the number of hydrogen atoms attached with a particular element.

	IA	IIA	IIIA	IVA	VA	VIA	VIIA
	NaH	MgH_2	AlH_3	SiH_4	PH_3	H_2S	H–Cl
Valency	1	2	3	4	3	2	1

Note: Valency w.r.t. H across the period increases up to 4 and then again decreases to 1.

Valency with respect to oxygen: Valency of 'O' = 2 It is defined as twice the number of oxygen atoms attached with a particular atom.

	IA	IIA	IIIA	IVA	VA	VIA	VIIA
	Na_2O	MgO	Al_2O_3	SiO_2	P_2O_5	SO_3	Cl_2O_7
Valency	1	2	3	4	5	6	7

Note : Valency with respect to oxygen increases from 1 to 7 across the period. Valency w.r.t. 'O' is equal to the group number.

New concept: This concept is based on the electronic configuration. According to this concept valency for IA to IVA group elements is equal to number of valence shell e^- and from VA to zero group, it is – [8– (number of valence e^-)].

Valency = No. of valence e^-

Valency = (8– no. of valence e^-)

	IA	IIA	IIIA	IVA	VA	VIA	VII	0
	ns^1	ns^2	ns^2np^1	ns^2np^2	ns^2np^3	ns^2np^4	ns^2np^5	ns^2np^6
Valence shell e^-	1	2	3	4	5	6	7	8
Valency	1	2	3	4	3	2	1	0
					(8 – 5) = 3			(8 – 8) = 0

Note : All the elements of a group have same valencies because they have same number of valence shell electrons.

Screening effect (σ) and effective nuclear charge (Z_{eff})

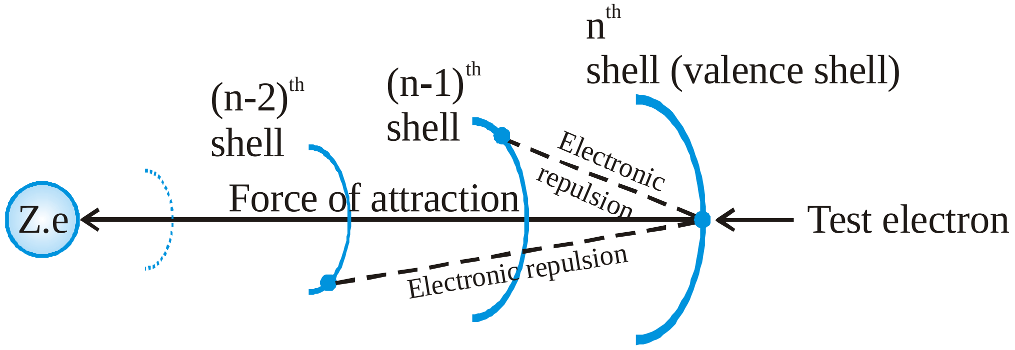
(a) Valence shell e^- suffer force of attraction due to nucleus and force of repulsion due to inner shell electrons.

(b) The decrease in force of attraction on valence e^- due to inner shell e^- is called screening effect or shielding effect.(i.e. total repulsive force is called shielding effect.)

(c) Due to screening effect. valence shell e⁻ experiences less force of attraction exerted by nucleus.(i.e. total attraction force experienced by valence e⁻ is called Zeff.)

(d) There is a reduction in nuclear charge due to screening effect. Reduced nuclear charge is called effective nuclear charge.

(e) If nuclear charge = Z, then effective nuclear charge = Z – σ (Where σ (Sigma)= Screening constant) So, Zeff = (Z – σ)



Slater's rule to know screening constant (σ)

(a) Screening effect (S.E.) of one e⁻ of the 1s is 0.30. **Ex.** In He ($1s^2$)

Screening effect of one 1s e⁻. where σ = 0.30

∴ Zeff = Z – σ = 2 – 0.30 = 1.7

(b) Screening effect of ns and np (Outermost orbit) electron is 0.35

(c) Screening effect of (n – 1) penultimate orbit s, p, d electrons is 0.85

(d) Screening effect of (n – 2) and below all the e⁻ present in s, p, d, f is 1.0

(Effective Nuclear charge of elements of second period)

Element	Electronic Configuration	Z	σ of ns & np electron	σ (n−1) orbital	Total Screening Constant (a + b)	Effective nuclear charge Z*=Z – σ
$_3$Li	$1s^2\ 2s^1$	3	–	0.85× 2=1.70	1.70	1.30
$_4$Be	$1s^2$, $2s^2$	4	1× 0.35=0.35	0.85× 2=1.70	2.05	1.95
$_5$B	$1s^2$,$2s^2$,$2p^1$	5	2× 0.35=0.70	0.85× 2=1.70	2.40	2.60
$_6$C	$1s^2$,$2s^2$,$2p^1$	6	3× 0.35=1.05	0.85× 2=1.70	2.75	3.25
$_7$N	$1s^2$,$2s^2$,$2p^3$	7	4× 0.35=1.40	0.85× 2=1.70	3.10	3.90
$_8$O	$1s^2$,$2s^2$,$2p^4$	8	5× 0.35=1.75	0.85× 2=1.70	3.45	4.55
$_9$F	$1s^2$,$2s^2$,$2p^5$	9	6× 0.35=2.10	0.85× 2=1.70	3.80	5.20

Periodic variation

(a) From left to right in a period Zeff increases

(I) That is why in a period Zeff increases by 0.65 and hence atomic size decreases considerably.

(II) In transition series Z increase by + 1 but screening effect increases by 0.85 So Zeff is 0.15

(1– 0.85 = 0.15) [Because e^- enters in (n – 1) orbit which has value of σ = 0.85]

In transition series Zeff increases very less amount, by 0.15 from left to right and hence atomic size remains almost constant.

Element	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn
Zeff	3.00	3.15	3.30	3.45	3.60	3.75	3.90	4.05	3.70	4.35

(b) From top to bottom in a group Zeff remain constant

Element	Li	Na	K	Rb	Cs	Fr
Zeff	1.30	2.20	2.20	2.20	2.20	2.20

ATOMIC RADIUS

The average distance of valence shell e^- from nucleus is called atomic radius. It is very difficult to measure the atomic radius because –

(I) The isolation of single atom is very difficult.

(II) There is no well-defined boundary for the atom.

(The probability of finding the e^- is 0 only at infinity). So, the more accurate definition of atomic radius is –

- Half the inter-nuclear distance(d) between two atoms in a homoatomic molecule is known as atomic radius.
- This inter-nuclear distance is also known as bond length. Inter-nuclear distance depends upon the type of bond by which two atoms combine.

Based on the chemical bonds, atomic radius is divided into four categories –

1. Covalent radius
2. Ionic radius
3. Metallic radius
4. Vander waal radius

1. Covalent radius

One half of the distance between the nuclei (internuclear distance) of two covalently bonded atoms in homoatomic molecule is called the covalent radius of that atom. The covalent bond must be single covalent bond. The covalent radius (r_A) of atom A in a molecule A_2 may be given as:

$$r_A = \frac{d_{A-A}}{2}$$

i.e. the distance between nuclei of two single covalently bonded atoms in a homoatomic molecule is equal to the sum of covalent radii of both the atoms

$$d_{A-A} = r_A + r_A$$

In a heteroditopic molecule AB where the electronegativity of atoms A and B are different, the experimental values of internuclear distance $d_{A\text{-}B}$ is less than the theoretical values ($r_A + r_B$).

According to Schomaker and Stevenson –

$$D_{A\text{-}B} = r_A + r_B - 0.09\, \Delta_x$$

Where Δ_x is the difference of electronegativities of the atoms A and B.

According to Pauling – If the electronegativities of the two atoms A and B are x_A and x_B respectively then

$$D_{A\text{-}B} = r_A + r_B - (C_1x_A - C_2x_B)$$

C_1 and C_2 are the Stevenson's coefficients for atoms A and B respectively.

2. Metallic Radius

Metal atoms are assumed to be closely packed spheres in the metallic crystal. These metal atom spheres are considered to touch one another in the crystal. One half of the internuclear distance between the two closest metal atoms in the metallic crystal is called metallic radius.

Metallic > Covalent radius

For example – Metallic radius and covalent radius of potassium are 2.3 Å and 2.03 Å respectively.

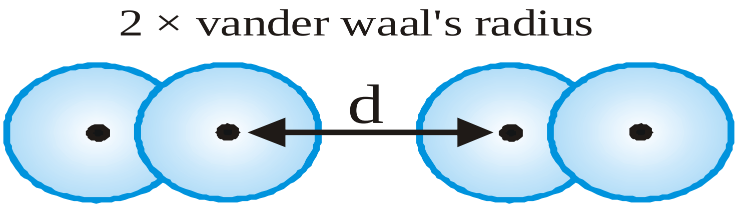
3. Van Der Wall's Radius or Collision radius

The molecules of nonmetal atoms are generally gases. On cooling, the gaseous state changes to solid state.

In the solid state, the nonmetallic elements usually exist as aggregations of molecules are held together by van der wall forces.

One half of the distance between the nuclei of two adjacent atoms belonging to two neighboring molecules of a compound in the solid state is called van der walls radius.

It may also be defined as half of the inter nuclear distance of two non-bonded neighboring atoms of two adjacent molecules.



Van der Wall's radius = $\frac{1}{2}$ × Internuclear distance between two successive nuclei of two covalent molecules (d)

Van der wall's radius > Metallic radius > Covalent radius

The van der walls radius and covalent radius of chlorine atom are 1.80 Å and 0.99 Å respectively

4. Ionic Radius

A neutral atom changes to a cation by the loss of one or more electrons and to an anion by the gain of one or more electrons. The magnitude of charge on cation and anion is equal to the number of electrons lost or gained respectively. The ionic radii of the ions present in an ionic crystal may be calculated from the inter molecular distance between the two ions.

(a) **Radius of Cation**

Radius of cation is smaller than that of corresponding atom.

Reasons

(I) During the formation of cation either one shell is removed or

(II) After removing an electron effective nuclear charge increase.

(b) **Radius of an Anion**

Radius of an anion is invariably bigger than that of the corresponding atom.

Reasons

(I) The effective nuclear charge decrease in the formation of anion. Thus the electrostatic force of attraction between the nucleus and the outer electrons decreases as the size of the anion increases.

(II) Interelectronic repulsion increases.

Factors affecting atomic radius are

(a) $\text{Atomic radius} \propto \dfrac{1}{\text{Effective nuclear charge (Zeff)}}$ $\text{Li} > \text{Be} > \text{B} > \text{C} > \text{N} > \text{O} > \text{F}$

(b)Atomic radius $\propto$ number of shells $\text{Li} < \text{Na} < \text{K} < \text{Rb} < \text{Cs}$

(c) Atomic radius $\propto$ Screening effect

(d) Atomic size $\propto$ Magnitude of –ve charge $< \text{O}^{-} < \text{O}^{-2}$

(e) Atomic radius $\propto$ $\text{Mn} > \text{Mn}^{+2} > \text{Mn}^{+3} > \text{Mn}^{+4}$

(f) Atomic radius $\propto \dfrac{1}{\text{Bond order}}$ $\quad >N-N< \; > -N\;\;N- \; > N\equiv N$

Periodic variation of atomic radius

(a) **Across a period:** It decreases from left to right in a period as nuclear charge increases

Ex. Li > Be > B > C > N > O > F < Ne

(b) **In a group:** It increases from top to bottom in a group as number of shell increases

Ex. Li < Na < K < Rb < Cs

Exceptions

(a) Transition elements

Sc Ti V Cr Mn

$Z_{eff} >$ Screening effect

Fe Co Ni

$Z_{eff} \approx$ Screening effect

(b) Lanthanide Contraction

(I) Outermost electronic configuration of inner transition elements is

$(n-2)\,f^{1-14}, (n-1)s^2p^6d^{0-1}, ns^2$ (n = 6 or 7)

(II) e^- enters in (n – 2) f orbitals

(III) Mutual screening effect of e^- is very less, because of complicated structure of f-orbital

(IV) Nuclear charge increases by one (+1) in lanthanides and actinides so atomic size of these elements slightly decreases. It is known as lanthanide contraction. Its effect is also observed in 5d transition series.

Here Nuclear charge > Screening effect.

(V) In Ist, 2nd and 3$^{rd.}$ transition series, Radii– 3d < 4d ≈ 5d (except IIIrd B)

	IIIB	IVB	
size	Sc	Ti ↓	size increases
increases	Y	Zr	
	La	Hf }	Equal due to lanthanide contraction

(c) Transition contraction

$\text{IIIA} \rightarrow \text{B} < \text{Al} \approx \text{Ga}$

Note : While atomic size should increases down the group.

(I) At. size of Ga = At. size of Al, due to transition contraction.

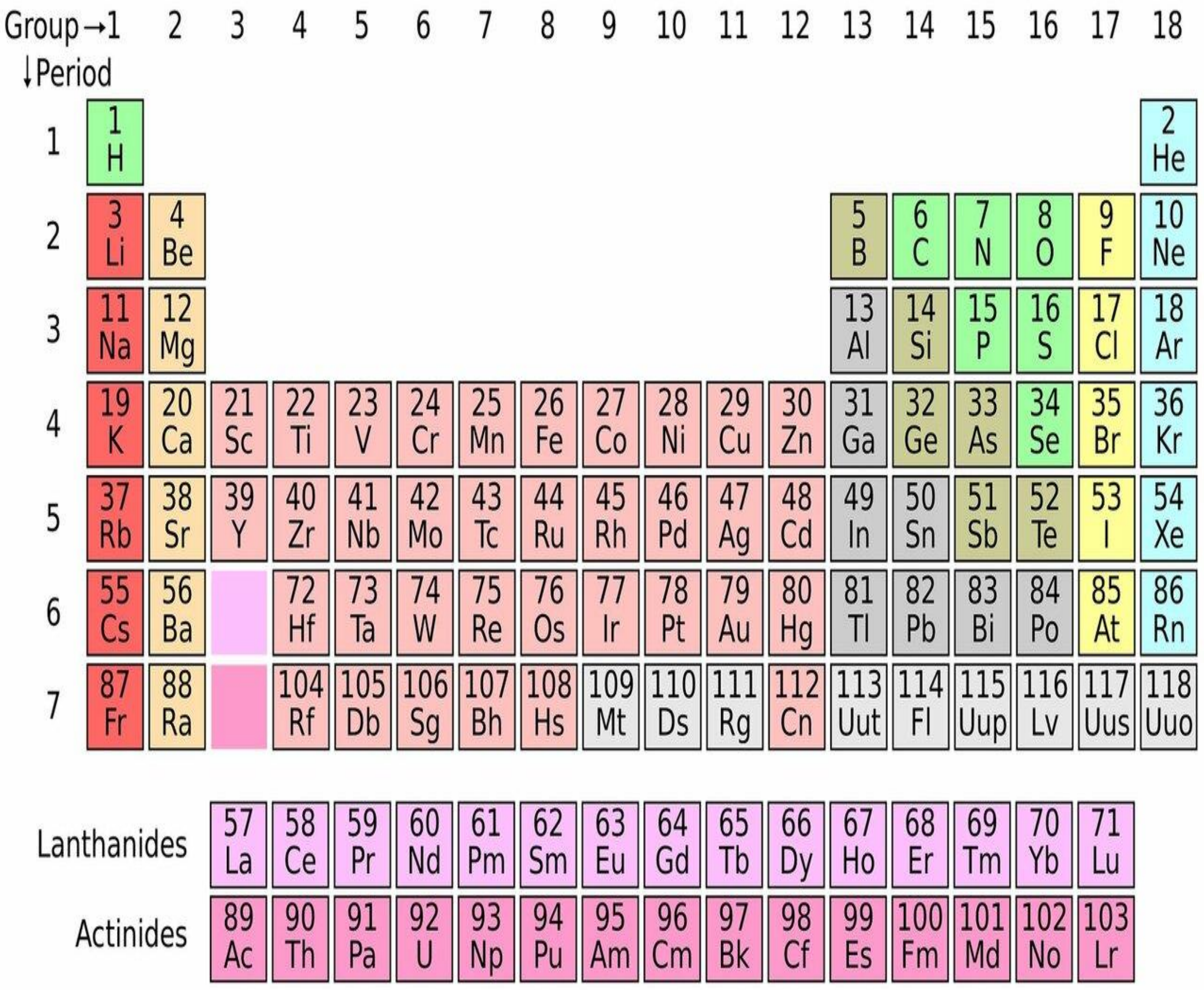
(II) In transition elements nuclear charge increases by 1.

(III) but e^- enters in $(n-1)d$ orbital exerts screening effect.

(IV) Screening effect of $(n-1)d$ e^- balance the nuclear charge by 85%

(V) Z_{eff} on increasing each electron = 1 – 0.85 = 0.15

(VI) Increase in nuclear charge is only 0.15 so atomic size remains almost constant.



Group→ ↓Period	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
5	37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
6	55 Cs	56 Ba		72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
7	87 Fr	88 Ra		104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Uut	114 Fl	115 Uup	116 Lv	117 Uus	118 Uuo
Lanthanides			57 La	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu	
Actinides			89 Ac	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr	

Covalent radius of the elements (In Å)

ISOELECTRONIC SERIES

A series of atoms, ions and molecules in which each species contains same number of electrons but different nuclear charge is called isoelectronic series.

	N^{3-}	O^{2-}	F^-	Ne	Na^+	Mg^{2+}
Number of e^-	10	10	10	10	10	10
Number of p	7	8	9	10	11	12

(a) Number of electrons is same.

(b) Number of protons is increasing.

(c) So the effective nuclear charge is increasing and atomic size is decreasing. In an isoelectronic series atomic size decreases with the increase of charge.

Some of the examples of isoelectronic series are as under.

$S^{2-}, Cl^-, K^+, Ca^{2+}, Sc^{3+}$ SO_2, NO_3^-, CO_3^{2-} N_2, CO, CN^- NH_3, H_3O^+

Ex. X – X bond length is 1.00 Å and C–C bond length is 1.54 Å. If electronegativities of X and C are 3.0 and 2.0 respectively, then C–X bond length is likely to be? (Using Stevenson & Schumacher formula).

Sol. $r_{C-X} = r_C + r_X - 0.09\,\Delta\chi$

$= -\ 0.09\ [\Delta\chi = 1] = 1.27 - 0.09\ ;$

Ans. C–X bond length = 1. 18 Å.

IONISATION POTENTIAL OR IONISATION ENERGY OR IONISATION ENTHALPY

Minimum energy required to remove most loosely held outer most shell e^- in ground state from an isolated gaseous atom is known as ionisation potential.

(Isolated → Without any bonding with other atom)

1. Successive Ionisation Energy

(a) For an atom M, successive ionisation energies are as follows -

$$M + E_1 \xrightarrow{\qquad} M^{+} + e^{-} \qquad E_1 = \text{Ionisation Potential}$$

$$M^{+} + E_2 \xrightarrow{\qquad} M^{+2} + e^{-} \qquad E_2 = \text{Ionisation Potential}$$

$$M^{+2} + E_3 \xrightarrow{\qquad} M^{+3} + e^{-} \qquad E_3 = \text{Ionisation Potential}$$

$$\text{Ionisation Potential} < \text{Ionisation Potential} < \text{Ionisation Potential}$$

(b) Electron cannot be removed from solid state of an atom, it has to convert in gaseous form, Energy required for conversion from solid state to gaseous state is called Sublimation energy.

(c) Ionisation Potential is always an endothermic process ($\Delta H = +ve$)

(d) It is measured in eV/atom (electron volt/atom) or Kcal/mole or KJ/mole

2. Factors affecting ionisation potential

(a) Atomic size: Larger the atomic size, smaller is the Ionisation Potential It is due to that the size of atom increases the outermost electrons e^{-} farther away from the nucleus and nucleus loses the attraction on that electrons and hence can be easily removed.

$$\boxed{\text{Ionisation Potential} \propto \frac{1}{\text{Atomic size}}}$$

(b) Effective nuclear charge (Zeff): Ionisation potential increases with the increase in nuclear charge between outermost electrons and nucleus.

$$\boxed{\text{Ionisation Potential} \propto \text{Effective nuclear charge}}$$

(c) Screening effect: Higher is the screening effect on the outer most electrons cause less attraction from the nucleus and can be easily removed, which is leading to the lower value of Ionisation Potential

$$\boxed{\text{Ionisation Potential} \propto \frac{1}{\text{Screening effect}}}$$

(d) Penetration power of sub shells

(I) Order of attraction of subshells towards nucleus (Penetration power) is -

$s > p > d > f$

(II) As subshell is more closer to nucleus so more energy will be required to remove e^- in comparison to p,d & f. **Ex.** Be B

$1s^2, 2s^2$ $1s^2, 2s^2\ 2p^1$

Ionisation Potential Be $>$ B

After losing one e^-, B attains electronic configuration of Be, so ionisation potential of B is more than

Be. II^{nd}onisation Potential of $B > Be$

(e) Stability of half-filled and fully filled orbitals

(I) Half-filled p^3,d^5, f^7 or fully filled s^2, p^6, d^{10}, f^{14} are more stable than others so it requires more energy.

Ex. N O

$1s^2, 2s^2\ 2p^3$ $1s^2, 2s^2\ 2p^4$ Ionisation Potential order is $O < N$

Because of half-filled p-orbitals in N, its ionisation energy (stability) is higher than O.

ionisation potential order $Na < Al < Mg$

(II) Because s-orbital in Mg is completely filled and its penetration power is also higher than p-orbital (Al). ionisation potential order

$$\underset{(2,8,1)}{Mg^+} < \underset{(2,8,2)}{Al^+} < \underset{(2,8)}{Na^+}$$

Periodic variation of ionisation energy

(a) Variation in period among the representative elements: Ionisation energies generally increases along the period because in moving left to right in a period the effective nuclear charge per outermost electron increases while the corresponding principal quantum number remain same.

(b)Variation in a group among the representative elements: The ionisation energy generally decreases in moving from top to bottom because the size increases due to the increase of the principal quantum number. On the other hand, the effective nuclear charge Zeff for the outermost electron remains almost the same along the group.

Application of ionisation potential

(a) Metallic and non-metallic character

Metallic ⟶ Ionisation Potential Low (Na, K, Rb etc.)

nonmetallic ⟶ Ionisation Potential High (F, Cl, Br etc.)

$$\boxed{\text{Ionisation Potential} \propto \frac{1}{\text{Metallic property}}}$$

(b) Reducing character

$$\boxed{\text{Reducing character} \propto \frac{1}{\text{Ionisation Potential}}}$$

(I) IA group has minimum ionisation potential so they are strong reducing agents in gaseous state ($Li < Na < K < Rb < Cs$)

(II) IA group – In Aqueous state reducing character $Li > K \simeq Rb > Cs > Na$

As the degree of hydration is more in Li due to high charge density.

(III) VIIA group has maximum ionisation potential so they are strong oxidising agents ($F > Cl > Br > I$)

(c) Stability of oxidation states

(I) If the difference between two successive ionisation potential $\geq$ 16eV then lower oxidation state is stable.

Ex. $Na \longrightarrow Na^{+}$ I

$$\left.\begin{array}{ll} Na \longrightarrow Na^{+} & I^{st} \text{ ionisation potential} \\ Na^{+} \longrightarrow Na^{+2} & II^{nd} \text{ ionisation potential} \end{array}\right\} 42.7 \text{ eV}$$

Difference of ionisation potential > 16 eV So Na^{+} is more stable.

(II) If the difference between two successive ionisation potential $\leq$ 11 then higher oxidation state is stable.

Ex.

$$\left.\begin{array}{ll} Mg \longrightarrow Mg^{+} & I^{st} \text{ ionisation potential} \\ Mg^{+} \longrightarrow Mg^{+2} & II^{nd} \text{ ionisation potential} \end{array}\right\} 7.4 \text{ eV}$$

Difference of ionisation potential $<$ 11 eV So Mg^{+2} is more stable.

$Al \longrightarrow Al^{+}$ } 12.8 eV So Al^{+} is more stable

$Al^{+} \longrightarrow Al^{+2}$ } 6.0 eV So Al^{+3} is more stable

$Al^{+2} \longrightarrow Al^{+3}$

Al^{+} is stable only in gaseous state

{

Al^{+3} is stable in liquid and solid state.

ELECTRON AFFINITY

(a) The amount of energy released when an electron is added to the outermost shell of one mole of an isolated gaseous atom in its lower energy state.

(b) The positive value of the electron affinity indicates that the process, i.e. $X_{(g)} + e^{-} \rightarrow$, is exothermic and the negative value indicates the process to be an endothermic one. Thus, the convention accepted in defining the electron affinity apparently contradicts the established convention in the thermodynamics.

Thus, $F_{(g)} + e^{-} \rightarrow$, $\Delta H = -328$ kJ mol^{-1}, and EA = 328 kJ mol^{-1} and

$N_{(g)} + e^{-} \rightarrow$, $\Delta H = +31$ kJ mol^{-1}, and EA = –31 kJ mol^{-1}

(c) Electron affinity just defined is actually first electron affinity since it corresponds to the addition of one electron only. In the process of adding further electron, the second electron will be added to gaseous anion against the electrostatic repulsion between the electron being added and the gaseous anion. Energy instead of being released is supplied for the addition of an electron to an anion.

$A_{(g)} + e^{-} \rightarrow A_{(g)}^{-} + E_1$; $A_{(g)}^{-} + e^{-} + \text{Energy supplied} \rightarrow A_{(g)}^{2-}$

Factors affecting electron affinity

(A) Atomic size $\quad \text{Electron Affinity} \propto \dfrac{1}{\text{Atomic size}}$

(B) Screening effect $\quad \text{Electron Affinity} \propto \dfrac{1}{\text{Screening effect}}$

(C) Effective nuclear charge (Z_{eff}) $\quad \text{Electron Affinity} \propto \text{Zeff}$

(d) Stability of completely filled or half-filled orbitals

Electron affinity of filled or half filled orbital is very less or zero. ⍰

Ex. (I) Elements (He, Ne, Ar , ...), Electron Affinity = Zero (⍰egH = + ve)

(II) Elements (Be, Mg, Ca, ...), Electron Affinity = $\simeq$ Zero (egH = + ve)

(III) Elements (N, P, As, ...), Electron Affinity = Very less

Periodic variation of electron affinity

(a) In a period: The effective nuclear charge increases and the size decreases with the increase of atomic number in a period. This is why, the electron affinity in general increases in a period. In fact, it reaches the climax for the Group VII A (i.e. halogens) elements.

Electron affinities (kJ mol^{-1})

IA	IIA	IIIA	IVA	VA	VIA	VIIA	Noble gas
Li 59.8	Be –60	B 27	C 122	N –31	O 141	F 328	Ne –90
Na 53	Mg –60	Al 44	Si 134	P 72	S 200	Cl 349	Ar 0
		Ga 29	Ge 120	As 77	Se 195	Br 324	

(b) **In a group:** For the representative elements, in moving down in a group generally the electron affinity falls down with the increases of atomic number because the effective nuclear charge Zeff at periphery per electron remain almost same but the size gradually increases due to addition of new shell.

ELECTRONEGATIVITY (EN)

(a) The tendency of an atom to attract shared electrons towards itself is called electronegativity.

(b) Electronegativity and Electron affinity both have tendency to attract electrons but electron affinity is for isolated atoms. Whereas electronegativity is for bonded atoms.

(c) A polar covalent or ionic bond of A – B may be broken as

(I) $A - B \longrightarrow A^{-}: + B^{+}$ (Electronegativity A > Electronegativity B)

or (II) $A - B \longrightarrow A^{+} + :B^{-}$ (Electronegativity A < Electronegativity B)

depending on their tendency to attract bonded electron.

(d) There is no unit of electronegativity because it is a relative value

(e) Electronegativity is property of a bonded atom not an isolated gaseous atom so no energy will be liberated or absorbed.

Difference between electronegativity and Electron Affinity

Electronegativity	Electron Affinity
Tendency of an atom in a molecule to attract the bonded electrons	Energy released when an electron is added to neutral isolated gaseous atom
Relative value of an atom	Absolute value of an atom
It regularly changed in a period or group	It does not changes regularly
It has no unit	It is measured in eV/atom or KJ mol^{-1} or K Cal $mole^{-1}$

Note: Small atoms are normally having more electronegativity than larger atoms.

Pauling's Scale

Pauling related the resonance energy (Δ_{AB}) of a molecule AB with the electronegativities of the atoms A and B. If x_A and x_B are the electronegativities of atoms A and B respectively then

$$0.208\sqrt{\Delta_{AB}} = x_A - x_B \text{ if } x_A > x_B$$

$$\text{or } \Delta_{AB} = 23.06\,(x_A - x_B)^2$$

$\Delta_{AB} = E_{AB(experimental)} - E_{AB(theoretical)}$ where E_{A-B} is the energy of A–B bond.

$$E_{AB\ (theoritical)} = \sqrt{(E_{A-A} \times E_{B-B})}$$

In a purely covalent molecule, AB, the experimental and theoritical values of bond energy A–B are equal,

$$\text{So } \Delta_{AB} = 0$$

$$\text{or } 0 = 23.06\,(x_A - x_B)^2$$

$$\text{or } x_A = x_B$$

In an ionic molecule AB, $E_{AB(experimental)}$ is more than $E_{AB(theoritical)}$

Pauling assumed the electronegativity value of fluorine 4 and calculated the electronegativity values of other elements from this value.

						H 2.1
Li 1.0	Be 1.5	B 2.0	C 2.5	N 3.0	O 3.5	F 4.0
Na 0.9	Mg 1.2	Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0
K 0.8	Ca 1.0	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8
Rb 0.8	Sr 1.0	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5
Cs 0.7	Ba 0.9	Tl 1.8	Pb 1.8	Bi 1.9	Po 2.0	
Fr 0.7	Ra 0.9					

In Pauling's scale, elements having almost same electronegativity are-

C = S = I = 2.5
N = Cl = 3.0
P = H = 2.1
Cs = Fr = 0.7
Be = Al = 1.5

Mullikan's Electronegativity

According to Mullikan, the electronegativity of an element is the average value of its ionisation potential and electron affinity.

$$\text{or} \quad \text{electro-negativity} = \frac{\text{Electron affinity + Ionisation potential}}{2}$$

When both are expressed in electron volt

Allred Rochon's Electronegativity: The electronegativity of an element is the electrostatic force of attraction between the electron present on the circumference of outermost shell of this atoms and the atomic nucleus. If the distance between the circumference of outermost shell & the nucleus is r and the effective nuclear charge Z_{eff} then

$$\text{Electron-negativity} \quad \frac{Z_{eff}\, e^2}{r^2} = \frac{0.359\, Z_{eff}}{r^2} = +0.744 \quad \Rightarrow \quad Z_{eff} = Z - \sigma$$

Z = The actual number of charge present in the nucleus i.e. number of protons and σ = Shielding constant.

Factors Affecting electronegativity

(A) Atomic size

$$\text{Electronegativity} \propto \frac{1}{\text{Atomic size}}$$

(B) Effective nuclear charge (Z_{eff})

$$\text{Electronegativity} \propto \text{Zeff}$$

(C) **Hybridisation state of an atom**

$sp > sp^2 > sp^3$

s character	50%	33%	25%
Electronegativity	3.25	2.75	2.5

(D) **Oxidation state**

$Mn^{+2} < Mn^{+4} < Mn^{+7}$

$O^{-2} < O^{-1} < O < O^{+1} < O^{+2}$

$Fe < Fe^{+2} < Fe^{+3}$

(I) s-orbital is nearer to nucleus so by increasing s-character in hybridization state, Electronegativity also increases.

(II) Electronegativity does not depend on filled or half-filled orbitals, because it is a tendency to attract bonded electron, not to gain or loss electron from outside.

Periodic Variation

(a) **In a period:** In moving from left to right in a period, the size gradually contracts and the effective nuclear charge increases. This is why, the electronegativity gradually increases with the increase of atomic number in a period up to the group VIIA elements.

(b) **In a group:** In the representative elements, in moving down in a group, the size increase while Zeff per electron at the periphery remains almost constant. This is why, the electronegativity generally falls in a group with the increase of atomic number.

Application of electronegativity

(a) **Metallic and non-metallic nature**

Low electronegativity $\longrightarrow$ Metals

High electronegativity $\longrightarrow$ Non-Metals

Metallic character increases down the group but decreases along a period.

(b) **Bond length**

$$\boxed{\Delta \text{EN} \propto \frac{1}{\text{Bond length}}}$$

Here Δ EN = difference in electronegativities of bonded atoms

Note: HF has minimum bond length as expected because of much difference in the electronegativities of H and F.

(c) **Acidic strength of hydrides**

$$NH_3 < H_2O < HF$$

In a particular period acidic strength of hydride is depends on electronegativity of M in M – H bond. Higher the electronegativity of M greater the acidic strength of hydride.

(d) **Nature of bonds**

(I) According to Hanny & Smith formula

Ionic % = $16\,(X_A - X_B) + 3.5\,(X_A - X_B)^2$

Here X_A = Electronegativity of A

X_B = Electronegativity of B

If $X_A - X_B \geq 2.1$ Ionic % $> 50\%$ i.e. Ionic bond

If $X_A - X_B \leq 2.1$ Ionic % $< 50\%$ i.e. covalent bond

(II) According to Gallis

$X_A - X_B \geq 1.7$ Ionic

$X_A - X_B \leq 1.7$ Covalent

If $X_A = X_B$; then A – B will be non-polar.

Ex. H—H, F—F

If $X_A > X_B$ and difference of electronegativities is small then

$A^{\delta-}$ —— $B^{\delta+}$ bond will be polar covalent

(e) Nature of hydroxides

(I) As per Gallis, In AOH if electronegativity of A is more than 1.7 (Non metal) then it is acidic in nature.

(II) If electronegativity of 'A' is less than 1.7 (metal) then AOH will be basic in nature

Ex.	NaOH	ClOH
X_A	0.9	3.0
Nature	Basic	Acidic

(III) If $X_A - X_0 \geq X_0 - X_H$ then AO bond will be more polar and will break up as

$A —— OH \longrightarrow A^+ + OH^-$ It shows basic nature

(IV) If $X_A - X_O \leq X_O - X_H$

$A — O —— H \longrightarrow H^+ + AO^-$ It shows Acidic nature

Ex. **In NaOH**

$X_O - X_{Na}$ (2.6) $>$ $X_O - X_H$ (1.4) So hydroxide is basic

In ClOH

$X_O - X_{Cl}$ (0.5) $<$ $X_O - X_H$ (1.4) So hydroxide is acidic

(f) Nature of oxides : Consider an oxide AO

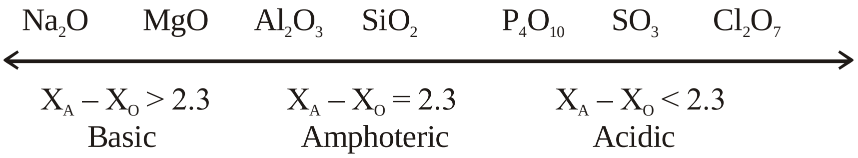
If $X_A - X_O > 2.3$ Basic oxide

If $X_A - X_O = 2.3$ Amphoteric oxide

If $X_A - X_O < 2.3$ Acidic oxide

(I) Along a period acidic nature of oxide increases.

(II) Down the group basic nature of oxide increases



ie. when in periodic table the distance between the element and oxygen increases, basic character increases.

$NO_2 > ZnO > K_2O$

⟶

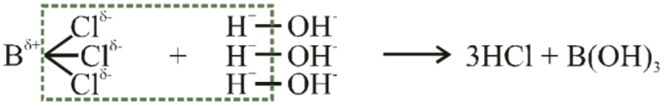
acidic character decreases

(g) Hydrolysis products of AX

Where A = Other element and X = Halogen

(I) If electronegativity of X > Electronegativity of A then on hydrolysis product will be HX.

Ex. (BCl_3), Electronegativity of Cl > Electronegativity of B



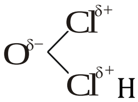
$PCl_3 + 3HOH \longrightarrow 3HCl + H_3PO_3$

$PCl_3 . Cl_2 + H_2O \longrightarrow 2HCl + POCl_3$

$POCl_3 + 3HOH \longrightarrow 3HCl + H_3PO_4$

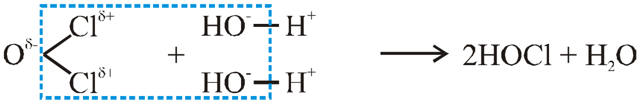
(II) If electronegativity of X < electronegativity of A then on hydrolysis product will be HOX (hypohalous acid)

Ex. Cl_2O



Here electronegativity of O > Cl

So On hydrolysis –



(h) Nomenclature of inorganic compounds

Prefix —— less electronegative element

Suffix —— More electronegative element

DENSITY

(a) In a group volume of an atom increases along with atomic weight but atomic weight increases more than atomic volume, So density increases in a group.

(b) Factors affecting density are

(I) Inner shell configuration

(II) Atomic weight

(III) Packing capacity

(IV) Chemical bond

Periodic Variation

(a) **In period** density first increases till maximum and then decreases. (s-block to d-block increases, d-block to p-block decreases)

(b) **In group** from top to bottom density increases regularly

Ex. In VIIA group – F and Cl are gases (Low density)

Br is liquid (density 3.19 g/cm^{-3})

I is solid (density 4.94 g/cm^{-3})

From s-block to p-block packing capacity and strength of bond increases.

BOILING POINTS AND MELTING POINTS

(a) It is a property of aggregate of atoms and not of a single atom.

i.e. why it is a molecular property.

(b) Melting point of a solid depend on –

(I) Structure of solid

(II) Chemical bond

(III) Bond energy

(c) Structure of solid made up –

(I) atomic solid

(II) molecular solid

(III) metallic solid.

Order of decreasing melting point is : Atomic solid > Metallic solid > Molecular solid

(d) Chemical bond among particles in solid. Order of bond strength is –

Covalent bond > Metallic bond > Vander waal bond

(e) Bond energy – Covalent solids like diamond, SiO_2 etc have only covalent bonds between atoms so their bond energy is higher than molecular or metallic solids.

Periodic variation

In period from left to right boiling point and melting point first increases then decreases.

Alkali metals – Crystal structure BCC (low boiling point & melting point)

Transition metals – ...."..... FCC (High boiling point & melting point)

Inert gases .... Lowest boiling point & melting point (Vander waal force)
Transition elements .... Highest melting point

Metals
W(Tungston) maximum melting point (3410° C)
Hg(Mercury) lowest melting point (– 38° C)

Non metals
Carbon (In the form of diamond) Highest melting point (3727° C)
Helium Lowest melting point (- 270° C)

In Group

(a) In s-block elements boiling point & melting point decreases down the group.

Li, Na (solid) → Cs, Fr (liquid)

It is due to weak cohesive energy of metallic bond in IA.

(b) In d-block elements boiling point & melting point increases down the group (due to lanthanide contraction, Zeff increases and hence bond energy increases)

(c) In p-block elements

(I) From IIIA – IVA group boiling point & melting point decreases down the group and from VA to '0' group, boiling point & melting point increases down the group. (Atomic or molecular weight $\propto$ vander waal force)

(II) boiling point and melting point of monoatomic molecules are lesser than diatomic molecules. '0' group < Halogens

(III) Atomic solid non-metals like B, C and Si has higher boiling point and melting point due to strong covalent bond.

(IV) Boiling point & melting point of molecular solids are less because of weaker vander waal force among molecules Ex. I_2.

Order of melting point

$Li > Na > K > Cs$

$Be > Ca > Sr > Ba > Mg$

$F_2 < Cl_2 < Br_2 < I_2$

$C > Si > Ge$

$C > B > Si$

$B > Al > Ga$

$He < Ne < Ar < Rn$

$Zn > Cd > Hg$

$Cu > Au > Ag$

BASIC INORGANIC NOMENCLATURE

Oxidation Number

- It is an imaginary or apparent charge developed over atom of an element when it goes from its elemental free state to combined state in molecules.
- It is calculated on basis of an arbitrary set of rules.
- It is a relative charge in a particular bonded state.
- In order to keep track of electron-shifts in chemical reactions involving formation of compounds, a more practical method of using oxidation number has been developed.

◆ In this method, it is always assumed that there is a complete transfer of electron from a less electronegative atom to a more electronegative atom.

Rules governing oxidation number

The following rules are helpful in calculating oxidation number of the elements in their different compounds. It is to be remembered that the basis of these rule is the electronegativity of the element

❖ Fluorine atom

Fluorine is most electronegative atom (known). It always has oxidation number equal to –1 in all its compounds

❖ Oxygen atom

In general, and as well as in its oxides, oxygen atom has oxidation number equal to –2.

In case of (I) peroxide (e.g. H_2O_2, , Na_2O_2) is –1,

(II) super oxide (e.g. KO_2) is –1/2

(III) ozonide (e.g. KO_3) is –1/3

(IV) in OF_2 is + 2 & in O_2F_2 is +1

❖ Hydrogen atom

In general, H atom has oxidation number equal to +1. But in metallic hydrides (e.g. NaH, KH), it is –1.

❖ Halogen atom

In general, all halogen atoms (Cl, Br, I) have oxidation number equal to –1. But if halogen atom is attached with a more electronegative atom than halogen atom, then it will show positive oxidation numbers.

Ex. $K\overset{+5}{Cl}O_3$, $H\overset{+5}{I}O_3$, $H\overset{+7}{Cl}O_4$, $K\overset{+5}{Br}O_3$

- **Metals**

  (a) Alkali metal (Li , Na, K, Rb, ...) always have oxidation number +1

  (b) Alkaline earth metal (Be , Mg , Ca ....) always have oxidation number +2.

  (c) Aluminium always has +3 oxidation number

  Note: Metal may have positive or zero oxidation number

- Oxidation number of an element in free state or in allotropic forms is always zero

  Ex. $\overset{0}{O_2}, \overset{0}{S_8}, \overset{0}{P_4}, \overset{0}{O_3}$

- Sum of the oxidation numbers of atoms of all elements in a molecule is zero.
- Sum of the oxidation numbers of atoms of all elements in an ion is equal to the charge on the ion
- If the group number of an element in modern periodic table is n, then its oxidation number may vary from (n – 10) to (n – 18) (but it is mainly applicable for p-block elements)

  Ex. N- atom belongs to 15th group in the periodic table, therefore as per rule, its oxidation number may vary from

  –3 to +5 ($N^{-3}H_3$, $N^{-2}O_3$, $N^{+4}O_2$, $N^{-5}{}_2 O_5$)

  Or ($\overset{-3}{N}H_3, \overset{+2}{N}O , \overset{+3}{N}_2O_3, \overset{+4}{N}O_2, \overset{+5}{N}_2O_5$)

- The maximum possible oxidation number of any element in a compound is never more than the number of electrons in valence shell. (But it is mainly applicable for p-block elements)

List of common oxidation state of an element in Periodic Table

1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1 H **+1** –1																	2 **He**
3 Li +1	3 Be +2											5 **B** **+3** –3	6 **C** **+4** +2 –4 etc.	7 **N** +5 +4 +3 +1 –3 **0** etc.	8 **O** +2 –1/2 –1 **–2**	9 **F** **–1**	10 **Ne**
11 Na +1	12 Mg +2											13 **Al** +3	14 **Si** +4 **–4**	15 **P** **+5** +3 +1 –3	16 **S** **+6** +4 +2 –2	17 **Cl** +5 +7 +3 +1 **–1** etc.	18 **Ar** **0**
19 **K** +1	20 **Ca** +2	21 **Sc** +2 **+3**	22 **Ti** +2 **+3** **+4**	23 **V** +2 **+3** +4 +5	24 **Cr** +2 **+3** +4 +5 +6	25 **Mn** **+2** +3 **+4** +5 +6 +7	26 **Fe** **+2** +3 +4 +5 +6	27 **Co** **+2** +3 +4 +5	28 **Ni** **+2** +3 +4	29 **Cu** **+2** +3	30 **Zn** **+2**	31 **Ga** +3	32 **Ge** +4 **–4**	33 **As** **+5** **+3** –3	34 **Se** **+6** +4 –2	35 **Br** +5 +3 +1 **–1**	36 **Kr** +4 +2 **0**
37 **Rb** +1	38 **Sr** +2											49 **In** **+3** +1	50 **Sn** **+4** +2	51 **Sb** **+5** **+3** –3	52 **Te** **+6** +4 –2	53 **I** +7 +5 +1 **0** **–1**	54 **Xe** +6 +4 +2 **0**
55 **Cs** +1	56 **Ba** +2											81 **Tl** +3 **+1**	82 **Pb** +4 **+2**	83 **Bi** +5 **+3**	84 **Po**	85 **At**	86 **Rn**

Elements

General Rule: The names of metals generally end in-ium or-um (examples are sodium, potassium, aluminum, and magnesium).

The exceptions are metals that were used and named in ancient times, such as iron, copper, and gold.

The names of nonmetals frequently end in-ine,-on, or -gen (such as iodine, argon, and oxygen.)

Given the names of the constituent elements and common ions, most of the common inorganic compounds can be named using the rules presented below.

Acids

Acids are normally classified in two groups, hydracids and oxyacid's

Hydracids

Hydracids are acids which contain hydrogen and a non-metal , but no oxygen.

General Rule: The names of hydracids have the prefix hydro-(sometimes shortened to hydr-) and the suffix-ic attached to the stem based on the names of the constituent elements (other than hydrogen.)

For example: HCl (made of hydrogen and chlorine) is hydrochloric acid; HBr (made of hydrogen and bromine) is hydrobromic acid; Hl (made of hydrogen and iodine) is hydroiodic acid; HCN (made of hydrogen, carbon, and nitrogen) is hydrocyanic acid; and H_2S (made of hydrogen and sulfur) is hydro sulfuric acid.

Oxoacids or Oxyacid's

The acids which contain hydrogen , oxygen and a metal or non-metal.

In this case , more than one possibilities arise because of the presence of different number of oxygen atoms. An example of such an oxoacid series is as follows: HClO, $HClO_2$, $HClO_3$, $HClO_4$. All these contains same three elements but differ in the number of oxygen atoms present.

General Rule - 1

If a class of acids contains only one member, its name is given the suffix-ic.

For example, hydrogen , carbon and oxygen combine to form only one acid i.e. H_2CO_3. It is called carbonic acid (carbonic acid.)

General Rule - 2

If an acid series contains two acids, such as H_2SO_4 and H_2SO_3, the acid containing more oxygen atoms is given the suffix -ic, while the acid with fewer oxygen atoms is given the suffix-ous.

For example : H_2SO_4 is sulfuric acid , and H_2SO_3 is sulfurous acid

Similarly, HNO_3 is nitric acid and HNO_2 is nitrous acid.

In the case of an extensive acid series (such as HClO, $HClO_2$, $HClO_3$, $HClO_4$), the acid with the one oxygen atoms lesser than -ous acid is given the prefix hypo- and the suffix -ous, and the acid with the one oxygen atom more than the -ic acid is given the prefix per and a suffix-ic.

In the above example, HClO is hypochlorous acid $HClO_2$ is chlorous acid, $HClO_3$ is chloric acid, and $HClO_4$ is perchloric acid.

A very valid doubt may arise, "How to decide which one is -ic acid to begin with ? This is somewhat arbitrary and has no unique answer. It depends upon the name given to it by old chemists, when no systematic nomenclature existed. Now these names have becomes so popular that it is difficult to change them. Do not worry, as you go through these notes, you will get a sufficient idea regarding these names. Cations (Positive ions)

Metal atoms with single positive charge

Rule: Names of positive ions end in-ium if the ion has only one oxidation state (Only one level of net charge). For example, the positive ion of sodium is Na^+ (sodium ion), and the positive ion of aluminum is Al^{3+} (aluminum ion).

Metal atoms with more than one possible charge

Rule: If the cation has variable valency (charge), charge is specified in roman numerals in round brackets immediately after the name of metal atom. For example, Sn^{2+} is written as tin (II) ion.

Alternately, the less positive ion ends in -ous, and the more positive ion ends in-ic . For instance, the two positive ions of copper are Cu^+ (cuprous) and Cu^{2+} (cupric). The oxidation state of a positive ion can also be designated by placing a Roman numeral after the name of the elements. These positive ions of copper can also be written as copper (I) and copper (II), respectively.

Ions	**Name**
Cu^{+}	cuprous ion
Sn^{2+}	Stannous ion
Sn^{4+}	Stannic ion
Fe^{3+}	Ferric ion
Fe^{2+}	Ferrous ion

General Rule-3

Suffix-nium is often used with cations containing non-metals.

For example, the positive ion of ammonia is NH_4^{+} (ammonium) and the positive ion of water (H_2O) is H_3O^{+} or H^{+} (hydronium).

Remember these names !

NO_2^{+} : nitronium **NO^{+} : nitrosonium** **H_3O^{+} : hydronium**

From NH_3 ammonia is derived NH_4^{+} : **ammonium.** Similarly.

N_2H_4 : hydrazine $\longrightarrow$ N_2H_5 : **hydrazinium**

$C_6H_5NH_2$: aniline $\longrightarrow$ $C_6H_5NH_3^{+}$:**anilinium**

C_5H_5N : pyridine $\longrightarrow$ $C_5H_5NH^{+}$: **pyridinium**

Anions (Negative Ions)

Anions can always be looked upon as ions derived from acids by removal of one or more protons. Accordingly, anions can be classified as follows.

Anions derived from hydracids

Rule : Names of negative ions from hydracids end in –ide.

For example , Cl^{-} (chloride) from HCl, and CN^{-} (cyanide) from HCN. Following examples will give you a better insight in this nomenclature. It is also useful to remember them.

Remember these names

F^-	fluoride	from HF	
Cl^-	chloride	from HCl	
Br^-	bromide	from HBr	
I^-	iodide	from HI	
O^{2-}	oxide	S^{2-}	sulphide
N^{3-}	nitride	P^{3-}	phosphide
As^{3-}	arsenide	Cu^-	cupride
H^-	hydride	Au^-	auride

Anions derived from oxyacids

Anion derived from an oxyacid by removal of one or more H^+ ions is termed as oxyanion.

Rule: If the oxyacid is – ic acid, suffix - ate is used with oxy-anion.

For example

CO_3^{2-}	carbonate (from H_2CO_3)
ZnO_2^{2-}	zincate
SiO_3^{2-}	silicate

Rule : If the oxyacid is - ous acid, suffix -ite is used with oxy-anion.

For example

NO_2^- (nitrite) is derived from HNO_2 (nitrous acid), and SO_3^{2-} (sulfite) is derived from H_2SO_3 (sulfurous acid)

Rule : If the oxyacid has prefixes per-or hypo–, the oxyanion will have same prefixes.

For example,

ClO_4^- percholrate ion from $HClO_4$, perchloric acid

ClO^- hypochlorite ion from HClO, hypochlorous acid

Remember these names !

SO_4^{2-} sulphate SO_3^{2-} Sulphite NO_3^{-} nitrate, NO_2^{-} nitrite

SnO_3^{2-}stannate SnO_2^{2-}stannite, PbO_3^{2-}plumbate, PbO_2^{2-}plumbite

Anions containing replacable hydrogen ions

Polyprotic acid. Any acid containing more than one replacable hydrogens is said to be a polyprotic acid.

Replacable hydrogens. H atoms which can be lost as H^{+} in reactions with a base.

H atoms connected to O atoms in oxyacid's are all replacable. If all the replacable hydrogens are removed, we obtain the anions discussed in the sections above.

However, in all the polyprotic acids it is always possible to remove less than the maximum number of replacable hydrogens . e.g. H_3PO_4 is triprotic. We can remove one, two or three H^{+} ions from it to generate $H_2PO_4^{-}$, , HPO_4^{2-} and PO_4^{3-} . You are already familiar with phosphate ion, PO_3^{-}. The other two anions, $H_2PO_4^{-}$ and HPO_4^{2-} still contain H atoms that are replacable. We consider their nomenclature in this section.

Rule-1 : A prefix bi- (old notation) or hydrogen – (IUPAC notation) is attached to the name of anion.

Rule -2 : For triprotic or higher acids, numerical prefixes (e.g. mono, bi, tri) are also used to indicate the number of replacable H atoms left in the sample).

Ex. HCO_3^{-} is bicarbonate or hydrogen carbonate

HSO_3^{-} bisulphite or hydrogen sulphite

HS^{-} bisulphide or hydrogen sulphide etc.

when anion has – 3 charge , e.g. PO_4^{3-} then following possibilities arise

HPO_4^{2-} monohydrogen phosphate, $H_2PO_4^{-}$ dihydrogen phosphate

Miscellaneous Anions (To be committed to memory)

NH_2^{-} amide, NH^{2-} imide, N^{3-} nitride, N_3^{-}azide, O_2^{2-} peroxide, O_2^{-} superoxide, O_3^{-} ozonide

CrO_4^{2-} chromate, $Cr_2O_7^{2-}$ dichromate, MnO_4^{2-} manganate, MnO_4^{-} permanganate , CN^{-} cyanide

OCN^{-} cyanate, SCN^{-} thiocyanate, $S_2O_3^{2-}$ thiosulphate, CH_3COO^{-} acetate, $C_2O_4^{2-}$ oxalate

$[Fe\ (CN)_6]^{3-}$ ferricyanide ,$[Fe(CN)_6]^{4-}$ ferrocyanide

Difference between Atoms and ions

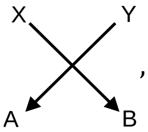
We have been talking for quite sometime. The distinction between atoms and ions, though simple, must be very clear in your mind. The following are the points summarise this difference.

Method of writing formula of an ionic compound

In order to write the formula of an ionic compound which is made up of two ions (simple or polyatomic) having net charges x and y respectively, follow the following procedure.

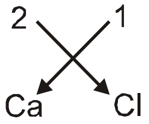
(I) Write the symbols of the ions side by side in such a way that positive ion is at the left and negative ion at the right as AB.

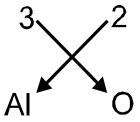
(II) write their charges on the top of each symbol as A^xB^y.

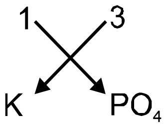
(III) Now apply crisis- cross rule as , i.e. formula A_yB_x.

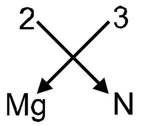
(IV) Cancel out any common factor (or HCF).

Ex. :

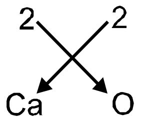
1. Calcium  chloride $= CaCl_2$

2. Aluminium oxide  $= Al_2O_3$

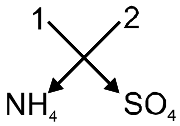
3. Potassium  phosphate $= K_3PO_4$

4. Magnesium 

nitride $= Mg_3N_2$

5. Calcium oxide 

$= Ca_2O_2$

6. Ammonium sulphate 

$= (NH_4)_2SO_4$

Cancelling the common factor, answer is CaO

	Atoms		Ions
1.	Atoms are perfectly neutral	1.	Ions are charged particles containing one or more atoms.
2.	In atoms, the number of portons is equal to the number of electrons. Na (protons 11, electrons 11) Cl (protons 17, electrons 17)	2.	In cations (positively charged ions), number of protons is more than the number of electrons. In anions (negatively charged ions) the no. of protons is less than the number of electrons. e..g. Na^+ (protons 11, electrons 10). Cl^- (protons 17, electrons 18)
3.	Except noble gases, atoms have less than 8 electrons in the outermost orbit e.g. Na 2, 8, 1; Ca 2, 8,8, 2 Cl 2, 8, 7 ; S 2, 8, 6	3.	Ions have generally 8 electrons in the outermost orbit , i.e., ns^3np^6 configuration. Na^+ , 2, 8; Cl^- 2, 8, 8 ; Ca^{2+} 2, 8, 8
4.	Chemical activity is due to loss or gain or sharing of electrons as to acquire noble gas configuration	4.	The chemical activity is due to the charge on the ion. Oppositely charged ions are held together by electrostatic forces.