Classification of elements
and
Periodic properties

Periodic table is arrangement of chemical elements in a tabular form on basis of their properties which facilitates the systematic study of properties of the elements.

Mendeleev's periodic Law : "The physical and chemical properties of elements are periodic functions of their atomic masses."

Drawbacks:

✓ No place for isotopes

✓ In few places, element having more atomic mass was placed before element having lower atomic mass.

✓ No unique place for H.

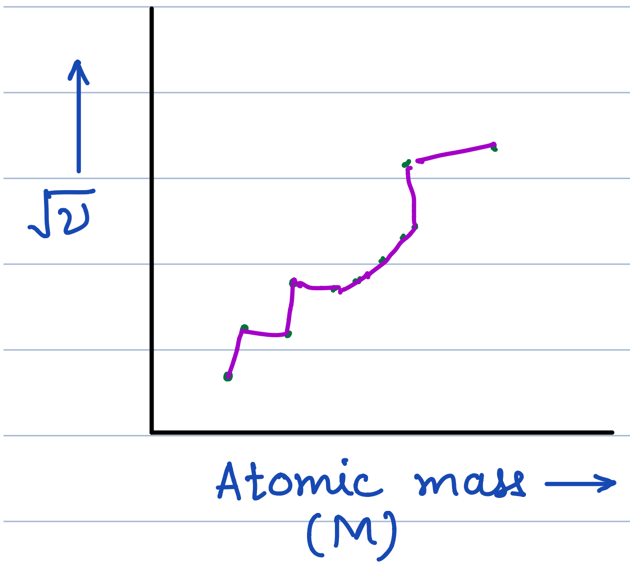
But Mendeleev's periodic table was challenged by Moseley's experiment.

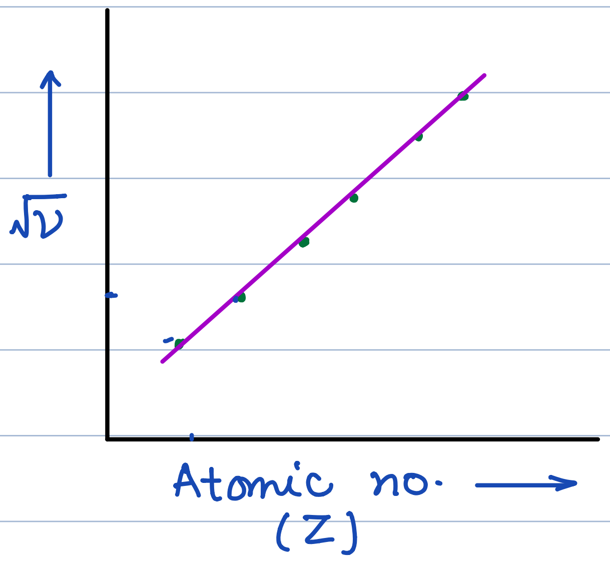
Moseley's experiment:

(Basis of modern periodic table)

He bombarded high speed (energetic) electrons on different heavy metal surfaces and obtained X-rays.

He observed a systematic mathematical relation between the frequency of emitted X-rays and atomic no. of metals used.





$\boxed{\sqrt{\nu} = a(Z-b)}$ Moseley's eq^n

ν = frequency of emitted X-rays
Z = atomic no. element
a, b = constants

$$\boxed{\sqrt{\nu} \propto Z}$$

He thereby showed that the atomic no. is a more fundamental property of an element than its atomic mass.

Derivation of Moseley's eq^n from Bohr's model

(not for school)

$$\nu = \frac{2\pi^2 m e^4}{h^3} Z^2 \left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right)$$

replacing Z by $\underbrace{Z_{effective}}_{\text{effective nuclear charge}}$ (e⁻–e⁻ repulsion is also considered)

$$\nu = \frac{2\pi^2 m e^4}{h^3} Z_{eff}^2 \left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right)$$

$$\nu = \frac{2\pi^2 m e^4}{h^3} (Z - b)^2 \left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right)$$

b → const. due to e⁻–e⁻ repulsion

✓ for K$_\alpha$- line $\boxed{\begin{matrix} n_f = 2 \\ n_i = 1 \end{matrix}} \rightarrow$ const.

$$\nu = \underbrace{\boxed{\frac{2\pi^2 m e^4}{h^3}\left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right)}}_{\text{const.} = a^2} (Z - b)^2$$

$$\boxed{\sqrt{\nu} = a(Z - b)}$$

Modern Periodic Law: The physical and chemical properties of elements are periodic functions of their atomic numbers.

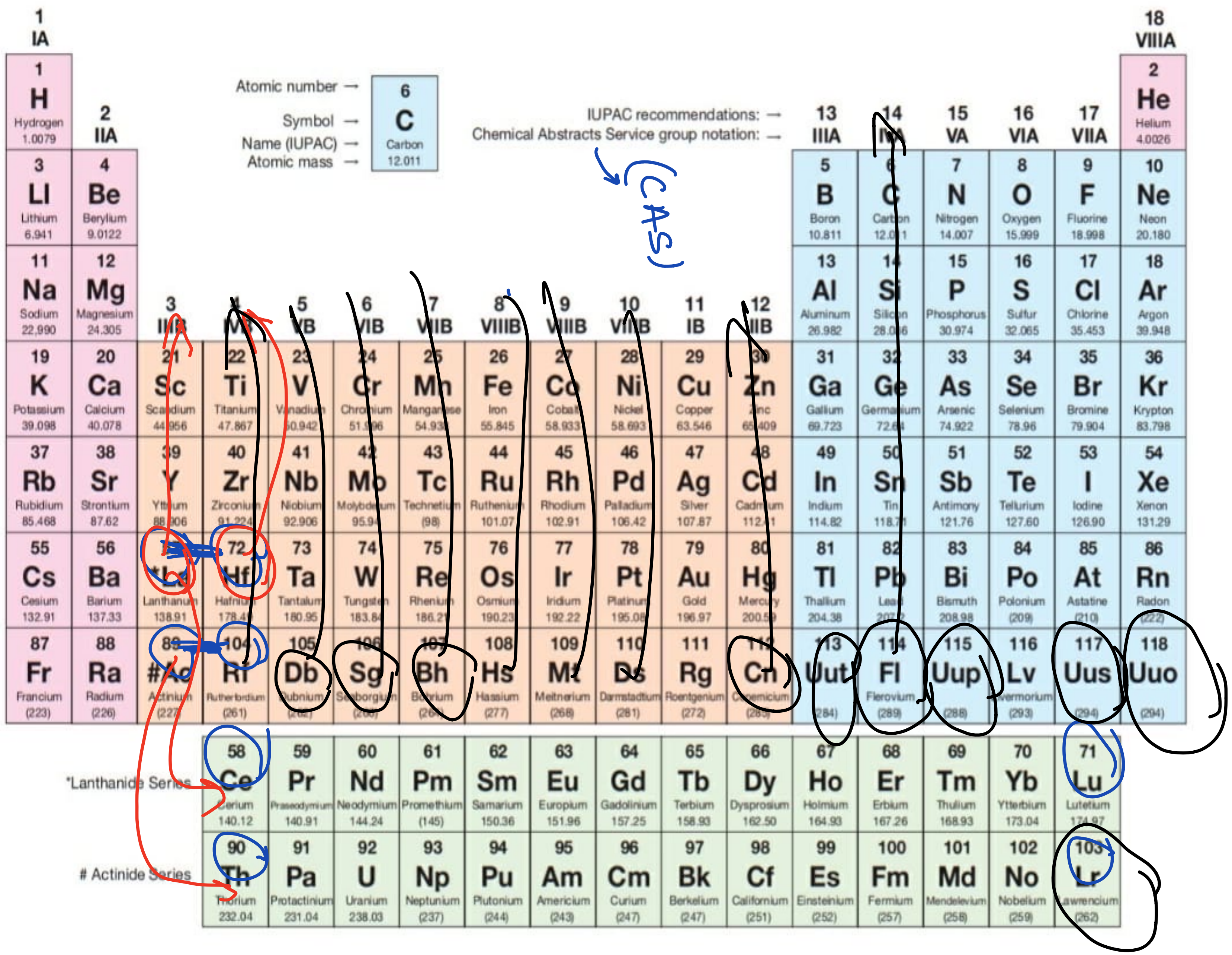
i.e. if elements are arranged in order of increasing atomic numbers, the elements with similar properties are repeated after certain regular intervals.

Modern periodic law is essentially the consequence of periodic variation in electronic configuration.

Long Form of modern periodic table:

- Total no. of known elements = 118
- Horizontal rows of periodic table are called "Periods" [Total no. of periods = 7]
- vertical columns of periodic table are called "Groups" (or) "families" [Total no. of Groups = 18]
- Total no. of blocks = 4 (s, p, d, f)

PERIODIC TABLE OF THE ELEMENTS



Atomic number → 6
Symbol → C
Name (IUPAC) → Carbon
Atomic mass → 12.011

IUPAC recommendations: →
Chemical Abstracts Service group notation: →

(CAS)

1 IA	2 IIA	3 IIIB	4 IVB	5 VB	6 VIB	7 VIIB	8 VIIIB	9 VIIIB	10 VIIIB	11 IB	12 IIB	13 IIIA	14 IVA	15 VA	16 VIA	17 VIIA	18 VIIIA
1 H Hydrogen 1.0079																	2 He Helium 4.0026
3 LI Lithium 6.941	4 Be Berylium 9.0122											5 B Boron 10.811	6 C Carbon 12.011	7 N Nitrogen 14.007	8 O Oxygen 15.999	9 F Fluorine 18.998	10 Ne Neon 20.180
11 Na Sodium 22.990	12 Mg Magnesium 24.305											13 Al Aluminum 26.982	14 Si Silicon 28.086	15 P Phosphorus 30.974	16 S Sulfur 32.065	17 Cl Chlorine 35.453	18 Ar Argon 39.948
19 K Potassium 39.098	20 Ca Calcium 40.078	21 Sc Scandium 44.956	22 Ti Titanium 47.867	23 V Vanadium 50.942	24 Cr Chromium 51.996	25 Mn Manganese 54.938	26 Fe Iron 55.845	27 Co Cobalt 58.933	28 Ni Nickel 58.693	29 Cu Copper 63.546	30 Zn Zinc 65.409	31 Ga Gallium 69.723	32 Ge Germanium 72.64	33 As Arsenic 74.922	34 Se Selenium 78.96	35 Br Bromine 79.904	36 Kr Krypton 83.798
37 Rb Rubidium 85.468	38 Sr Strontium 87.62	39 Y Yttrium 88.906	40 Zr Zirconium 91.224	41 Nb Niobium 92.906	42 Mo Molybdenum 95.94	43 Tc Technetium (98)	44 Ru Ruthenium 101.07	45 Rh Rhodium 102.91	46 Pd Palladium 106.42	47 Ag Silver 107.87	48 Cd Cadmium 112.41	49 In Indium 114.82	50 Sn Tin 118.71	51 Sb Antimony 121.76	52 Te Tellurium 127.60	53 I Iodine 126.90	54 Xe Xenon 131.29
55 Cs Cesium 132.91	56 Ba Barium 137.33	57 *La Lanthanum 138.91	72 Hf Hafnium 178.49	73 Ta Tantalum 180.95	74 W Tungsten 183.84	75 Re Rhenium 186.21	76 Os Osmium 190.23	77 Ir Iridium 192.22	78 Pt Platinum 195.08	79 Au Gold 196.97	80 Hg Mercury 200.59	81 Tl Thallium 204.38	82 Pb Lead 207.2	83 Bi Bismuth 208.98	84 Po Polonium (209)	85 At Astatine (210)	86 Rn Radon (222)
87 Fr Francium (223)	88 Ra Radium (226)	89 #Ac Actinium (227)	104 Rf Rutherfordium (261)	105 Db Dubnium (262)	106 Sg Seaborgium (266)	107 Bh Bohrium (264)	108 Hs Hassium (277)	109 Mt Meitnerium (268)	110 Ds Darmstadtium (281)	111 Rg Roentgenium (272)	112 Cn Copernicium (285)	113 Uut (284)	114 Fl Flerovium (289)	115 Uup (288)	116 Lv Livermorium (293)	117 Uus (294)	118 Uuo (294)

*Lanthanide Series	58 Ce Cerium 140.12	59 Pr Praseodymium 140.91	60 Nd Neodymium 144.24	61 Pm Promethium (145)	62 Sm Samarium 150.36	63 Eu Europium 151.96	64 Gd Gadolinium 157.25	65 Tb Terbium 158.93	66 Dy Dysprosium 162.50	67 Ho Holmium 164.93	68 Er Erbium 167.26	69 Tm Thulium 168.93	70 Yb Ytterbium 173.04	71 Lu Lutetium 174.97
# Actinide Series	90 Th Thorium 232.04	91 Pa Protactinium 231.04	92 U Uranium 238.03	93 Np Neptunium (237)	94 Pu Plutonium (244)	95 Am Americium (243)	96 Cm Curium (247)	97 Bk Berkelium (247)	98 Cf Californium (251)	99 Es Einsteinium (252)	100 Fm Fermium (257)	101 Md Mendelevium (258)	102 No Nobelium (259)	103 Lr Lawrencium (262)

No. of elements in periods:

Period No.	No. of elements
1	2
2	8
3	8
4	18
5	18
6	32
7	32

Formula:

$$= 2\left\{ Int\left(\frac{N+2}{2}\right)\right\}^2$$

N = period no.

Int(x) : round of integer of x

✓ N=1

$$= 2\left\{ Int\left(\boxed{\frac{1+2}{2}}\right)\right\}^2$$

$$= 2(1)^2 = 2$$

N=2

$$= 2\left\{ Int\left(\frac{2+2}{2}\right)\right\}^2$$

$$= 2\{2\}^2$$

$$= 8$$

Blockwise brif introduction of periodic table:

[A] s-block : (all metals)

✓ contains Group-1 & Group-2 metals

✓ General configuration : [noble gas] $ns^{1 \text{ or } 2}$

- Group-1 metals : known as alkali metals

∵ they form alkaline hydroxide on rxn with H_2O.

- Group-2 metals: Known as alkaline earth metals (except Be)

∵ They form alkaline oxides and hydroxides and their oxides are abundant in earth's crust.

BeO or $Be(OH)_2$: Amphoteric

so, Be is not included in alkaline earth metal.

[B] p-Block

- Contains metals, non-metals, metalloids & noble gases.
- Group no. 13 to Group no. 18 → p-block
- General configuration = [noble gas] $ns^2\ np^{1\text{ to }6}$

pnictogens	chalcogens	halogens
(Gr-15 element)	(Gr-16 elements)	(Gr-17 elements)
N, P, As, Sb, Bi	(O, S, Se, Te, Po)	(F, Cl, Br, I, At)
→ poisonous	→ ore forming	→ salt forming

Noble gases / inert gases:

- Group-18 elements (He, Ne, Ar, Kr, Xe ---)
- Least reactive

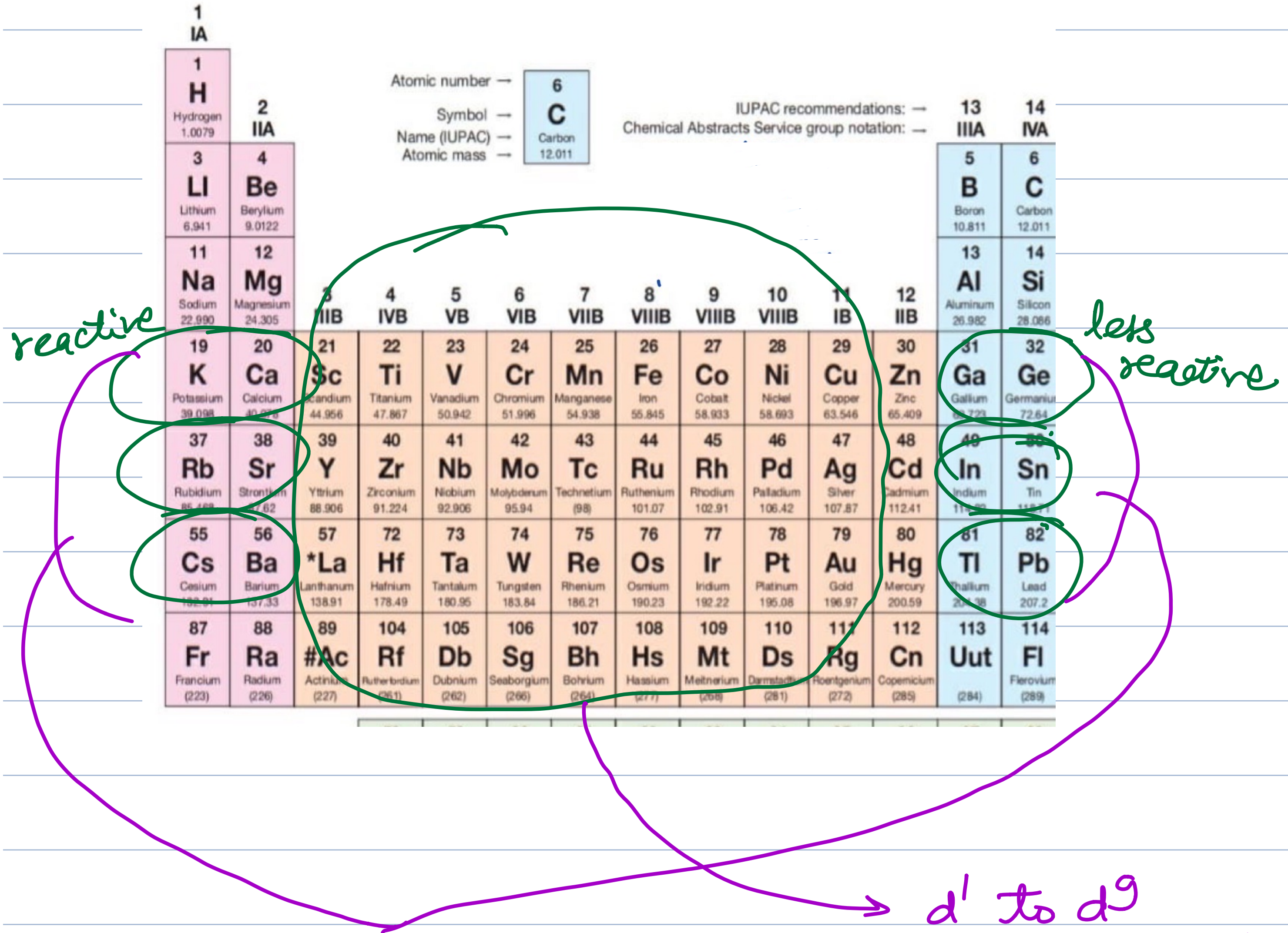
[C]. d- Block :

- ✓ all metals
- ✓ contains → Group-3 to Gr-12
- ✓ General configuration : [noble gas] $ns^{1\text{ to }2}\ (n-1)d^{1\text{ to }10}$

- All d-block metals except Group no. 12 (Zn, Cd, Hg) are called transition elements or

Transition metals.

Transition elements form a bridge between reactive metals of s-block and less reactive metals of p-block (i.e. Group-13).



Atomic number → 6
Symbol → C
Name (IUPAC) → Carbon
Atomic mass → 12.011

IUPAC recommendations: →
Chemical Abstracts Service group notation: →

1 IA	2 IIA	3 IIIB	4 IVB	5 VB	6 VIB	7 VIIB	8 VIIIB	9 VIIIB	10 VIIIB	11 IB	12 IIB	13 IIIA	14 IVA
1 H Hydrogen 1.0079													
3 Li Lithium 6.941	4 Be Beryllium 9.0122											5 B Boron 10.811	6 C Carbon 12.011
11 Na Sodium 22.990	12 Mg Magnesium 24.305											13 Al Aluminum 26.982	14 Si Silicon 28.086
19 K Potassium 39.098	20 Ca Calcium 40.078	21 Sc Scandium 44.956	22 Ti Titanium 47.867	23 V Vanadium 50.942	24 Cr Chromium 51.996	25 Mn Manganese 54.938	26 Fe Iron 55.845	27 Co Cobalt 58.933	28 Ni Nickel 58.693	29 Cu Copper 63.546	30 Zn Zinc 65.409	31 Ga Gallium 69.723	32 Ge Germanium 72.64
37 Rb Rubidium 85.468	38 Sr Strontium 87.62	39 Y Yttrium 88.906	40 Zr Zirconium 91.224	41 Nb Niobium 92.906	42 Mo Molybdenum 95.94	43 Tc Technetium (98)	44 Ru Ruthenium 101.07	45 Rh Rhodium 102.91	46 Pd Palladium 106.42	47 Ag Silver 107.87	48 Cd Cadmium 112.41	49 In Indium 114.82	50 Sn Tin 118.71
55 Cs Cesium 132.91	56 Ba Barium 137.33	57 *La Lanthanum 138.91	72 Hf Hafnium 178.49	73 Ta Tantalum 180.95	74 W Tungsten 183.84	75 Re Rhenium 186.21	76 Os Osmium 190.23	77 Ir Iridium 192.22	78 Pt Platinum 195.08	79 Au Gold 196.97	80 Hg Mercury 200.59	81 Tl Thallium 204.38	82 Pb Lead 207.2
87 Fr Francium (223)	88 Ra Radium (226)	89 #Ac Actinium (227)	104 Rf Rutherfordium (261)	105 Db Dubnium (262)	106 Sg Seaborgium (266)	107 Bh Bohrium (264)	108 Hs Hassium (277)	109 Mt Meitnerium (268)	110 Ds Darmstadtium (281)	111 Rg Roentgenium (272)	112 Cn Copernicium (285)	113 Uut (284)	114 Fl Flerovium (289)

reactive

less reactive

d^1 to d^9 configuration in atom or ions

✓ either d^0 or d^{10} configuration in atom or ion

Zn, Cd, Hg $\xrightarrow[\text{or ion}]{\text{either atom}}$ all have d^{10}

→ not included in transition element.

- In atom or ion of transition elements, they have partially filled d-subshell.

Ⓠ Cu has configuration [Ar] $4s^1\ 3d^{10}$. still it is included in transition elements. How?

Sol: Cu^{2+} (ion) : [Ar] $4s^0\ 3d^9$

✓✓ Zn, Cd, Hg $\Rightarrow$

Zn : [Ar] $4s^2\ 3d^{10}$ → fully filled d

Zn^{2+} : [Ar] $4s^0\ 3d^{10}$ → fully filled d

[D] f-block :

✓ all metals

✓ Group-3
- → at. no. 58 to 71 : Lanthanoids or Lanthanides
- → at. no. 90 to 103 : Actinoids or Actinides

✓ General configuration : [noble gas] $ns^2 (n-1)d^{0 \text{ to } 2} (n-2)f^{1 \text{ to } 14}$

- They are called inner transition elements. (at. no. in b/w transition elements)

Lanthanoids : 14 elements lying after lanthanum
Actinoids : 14 elements lying after actinium

Transurenic/ transurenium elements:

- Elements after at. no. 92 (Uranium)
- They are man-made (synthetic) not known in nature.

Representative elements/ Main group elements:

- Elements of s-block and p-block are collectively known as main group elements.

Finding group no., Block, & period no. for an element (whose atomic no. is given) in Periodic table:

Method '1' (for school exams):

Step: 1. write electronic configuration

Step:2. max. principal quantum no. = period no.

Step:3. subshell of last e^- = Block

Step:4. To find group no, we use following table:

Block	Group no.
s	= no. of $\bar{e}$s in outermost s
p	= no. of $\bar{e}$s in outermost p + 12
d	= no. of $\bar{e}$s in outermost s + no. of e^- in inner d (ns) (n-1) d
f	= 3

ex: atomic no. 15

$= 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^3$

Period = 3rd
Block = p
Group no. = (3+12) = 15

ex at. no. 24

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ \underbrace{4s^1}_{ns}\ \underbrace{3d^5}_{(n-1)d}$

Period = 4th
Block = d
Group no. = (1+5) = 6

---

atomic no. 12

$1s^2\ 2s^2\ 2p^6\ 3s^2$

Period = 3rd
Block = s
Group no. = 2

atomic no. 62

$\overset{54}{[Xe]}\ 6s^2\ 4f^6$

Period = 6th
Block = f
Group no. = 3

* It is failed in some cases:

Pd(46): $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 4d^{10}$

(actual configuration)

Period acc. to configuration: 4th
But actually its period is: 5th

Method '2':

s-block ⇒ Gr 1 & 2
p- " ⇒ Gr 13 to Gr 18
d- " ⇒ Gr 3 to Gr 12
f-block ⇒ at. no. 58 to 71
at. no. 90 to 103

Gr. no. 18	Period
He(2)	1st
Ne(10)	2nd
Ar(18)	3rd
Kr(36)	4th
Xe(54)	5th
Rn(86)	6th

ex at. no. 15

Period: 3rd
Gr: no. 15
p-block

at. no. 53

Period : 5th
Gr. no. : 17
p-block

- atomic no. from 103 to 118:

[Group. no. = atomic No. − 100]

Prediction of atomic no. of a given location of element in periodic table:

ex (i) [Period 2^{nd} Group. no. 16]

atomic No. = 8

ex (ii) [Period : 4^{th} Group: no. 5]

at. no = 23

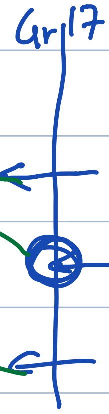
Gr: 18	Period ↓
He (2)	1st
Ne (10)	2nd
Ar (18)	3rd
Kr (36)	4th
Xe (54)	5th
Rn (86)	6th

Prediction of atomic no. of elements lying above and below for an element of given atomic no:

ex (i) [at. no. 17]

above element : at. no. 9

below element : at no. 35



Gr: 18	Period ↓
He (2)	1st
Ne (10)	2nd
Ar (18)	3rd
Kr (36)	4th
Xe (54)	5th
Rn (86)	6th

IUPAC naming of elements having atomic no. more than 100 (naming of superheavy elements):

- ✓ It is a temporary naming of elements.
- ✓ It is based on digits of atomic numbers of elements.

Digit	Root	abbreviation
0	nil	n
1	un	u
2	bi	b
3	tri	t
4	~~tetra~~ quad	q
5	pent	p
6	hex	h
7	~~hept~~ sept	s
8	oct	o
9	enn	e

- ✓ The common suffix is : "ium"
- ✓ Add the roots in sequence of digits shown and then add common suffix for naming.
- ✓ i + i = i

ex: at. no. 107

Un nil sept

Name: Unnilseptium

symbol: Uns

ex: at. no. 113

un un tri

Name: $\underbrace{\text{Ununtriium}}$

→ Ununtrium

Q Find group no. of element Unq

Sol: at. no. = 104

Group. no. = 104 − 100 = 4

Q. what can be group. no. of a hypothetical element whose atomic no. is 120 ?

Sol: Group-2 (s-block)

↳ alkaline earth metal

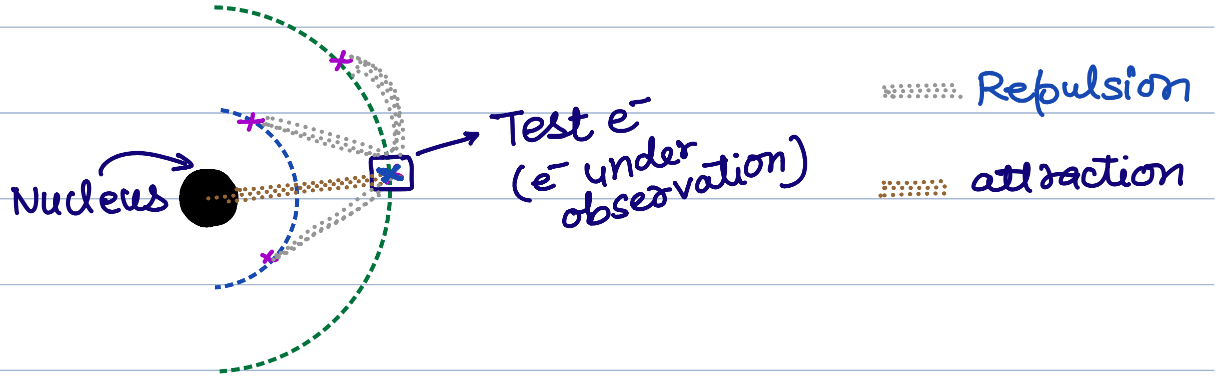
Some important effects:

[A] Penetration effect:

- ✓ Penetration describes the degree of closeness of an orbital w.r.t. the nucleus.
- ✓ For a given shell, penetration follows the order: $s > p > d > f$

[B]. Screening effect (Shielding effect):

Imp :



- Due to presence of inner electrons or other electrons present in same shell, the outer e^- does not experiance the full positive charge of the nucleus. (Full positive charge of nucleus = Ze)

- The effect is lowered due to the partial screening of positive charge of the nucleus by inner electrons.

- This is called shielding of outer electron from nucleus by the inner electrons or other electrons present in the same shell.

Element	configuration	screening to outermost e^- by
Li (3)	$1s^2\ 2s^1$	Two $\bar{e}$s of 1s
Be (4)	$1s^2\ 2s^2$	Two $\bar{e}$s of 1s + one left $\bar{e}$ of 2s
S (16)	$1s^2 2s^2 2p^6 3s^2 3p^4$	Two $\bar{e}$s of 1s + Two $\bar{e}$s of 2s + Six $\bar{e}$s of 2p + Two $\bar{e}$s of 3s + three left $\bar{e}$s of 3p

* The order of shielding effect exerted by electrons of a given shell on a test e^- is given as: $ns > np > nd > nf$ (n= shell no.)

- within a shell, s-orbital electrons spends more time close to the nucleus in comparison to p-orbital electrons which spends more time in the vicinity of nucleus as compared to d-orbital electrons.

[C]. Effective nuclear charge (Zeffective):

The net positive charge (i.e. actual charge that electron feels) of nucleus experienced by an electron in an atom is called effective nuclear charge.

$$\boxed{Z_{effective} = \underset{\substack{\downarrow \\ (\text{atomic no.})}}{Z} - \sigma}$$

σ — Screening const or Shielding const

For any single e^- species,

$$\left\{\begin{array}{l} \sigma = 0 \text{ (No other } e^- \text{ to shield)} \\ Z_{effective} = Z \end{array}\right.$$

***

[Across the period, effective nuclear charge increases from left to right.]

2nd Period :

Li	Be	B	C	N	O	F	Ne
$1s^2 2s^1$	$1s^2 2s^2$	$1s^2 2s^2$ $2p^1$	$1s^2 2s^2$ $2p^2$	$1s^2 2s^2$ $2p^3$	$1s^2 2s^2$ $2p^4$	$1s^2 2s^2$ $2p^5$	$1s^2 2s^2$ $2p^6$

inc. proton by 1 unit (Z inc. by 1 unit)
inc. electron by 1 unit (σ inc. but by less than 1 unit)

✓ Z inc. more rapidly than σ
hence, Z_{eff} inc.

Periodic properties:

(A) Atomic radius
(B) Ionic radius
(C) Ionisation energy
(D) Electron affinity
(E) Electronegativity

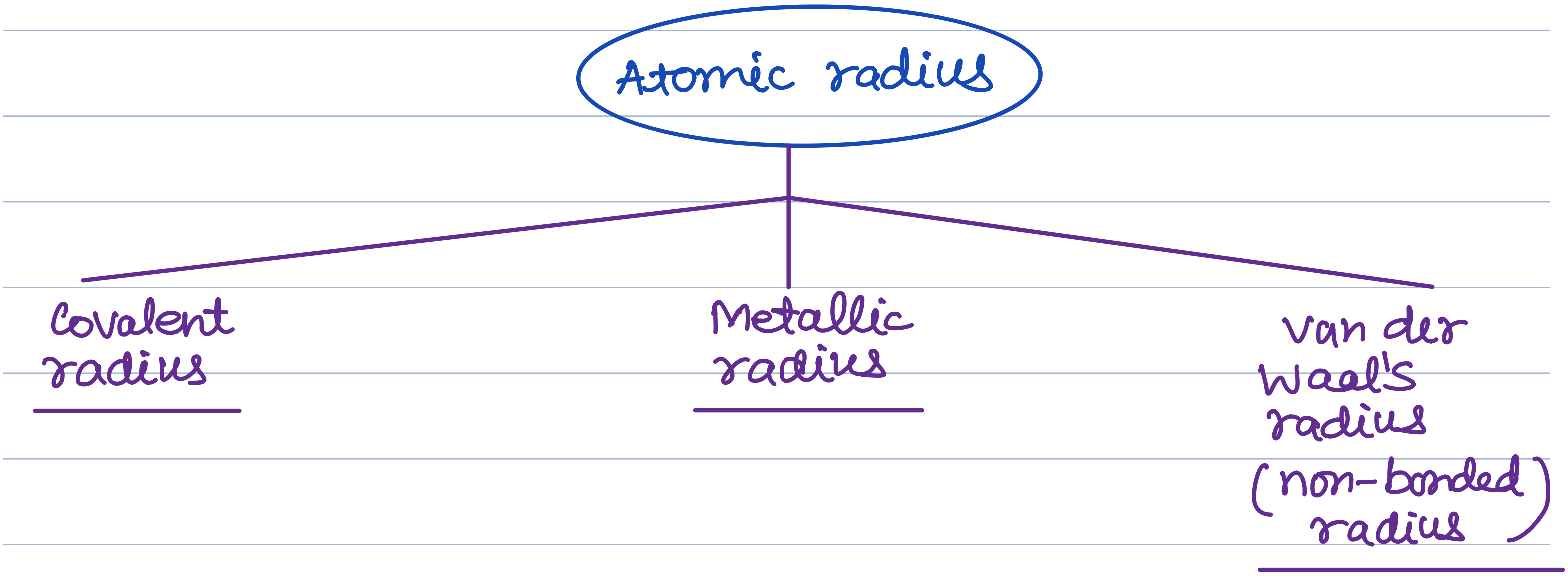
[1]. Atomic radius:

In an atom, the distance between nucleus and outermost electron cloud is atomic radius.

- Exact atomic radius of an atom can not be obtained as:

(i) size of atoms are too small (~ 1.2 Å)
(ii) there is no sharp boundary of electron cloud.
(iii) Atoms are mostly found in bonded form.
(iv) surroundings affects size

- size of atoms are studied in their respective bonded states.

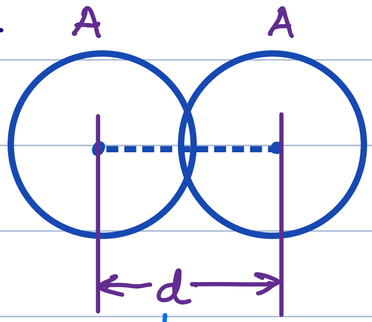


Covalent radius:
(SBCR : Single bond covalent radius)

Half of the internuclear distance in a single covalently bonded homonuclear molecule is covalent radius.

Internuclear distance: distance b/w two nucleus of bonded atoms.
Homonuclear molecule: Molecule made of same element

Homonuclear Molecule →



bond length (internuclear distance)

$$\boxed{(r_c) = \frac{d}{2}}$$

(r_c) → covalent radius of atom 'A'

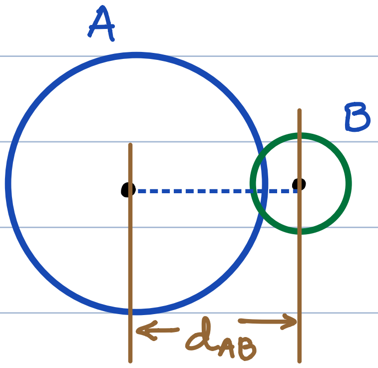
ex: Bond length in Cl_2 molecule = 198 pm

covalent radius of Cl atom = $\frac{198}{2}$ = 99 pm

✓ On Increasing bond multiplicity, internuclear distance decreases.

ex:	C–C	C=C	C≡C
d =	1.54 Å	1.33 Å	1.20 Å

Heteronuclear molecule:



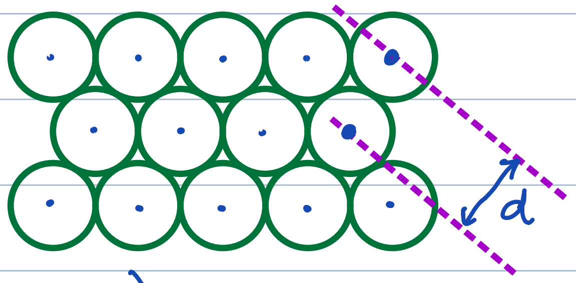
✓ Stevenson– schomaker's eqⁿ:

$$\boxed{d_{A-B} = (r_c)_A + (r_c)_B - 0.09\left|\chi_A - \chi_B\right|}$$

d_{A-B} = A–B bond length in Å
$(r_c)_A$ = covalent radius of A in Å
$(r_c)_B$ = covalent radius of B in Å
χ_A = Electronegativity of A
χ_B = electronegativity of B

Metallic radius :

Half of the internuclear distance of two closest atoms in a metallic crystal is called metallic radius.



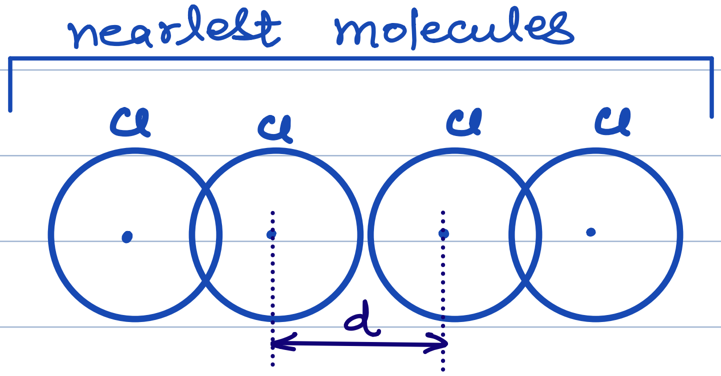
(Metallic lattice)

$$\boxed{r_m = \frac{d}{2}}$$

van der Waal's radius:

In solid state, half of the internuclear distance of two closest atoms belonging to the nearest molecules is called van der Waal's radius.

ex: ✓



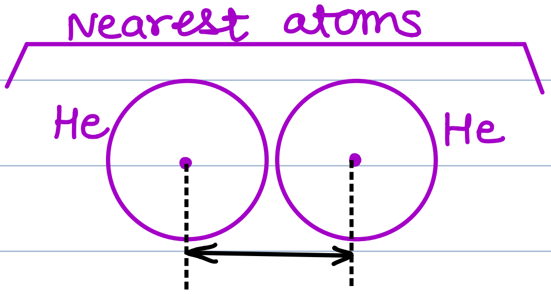
Cl_2 (solid state)

$$\boxed{(r_v)_{Cl} = \frac{d}{2}}$$

↳ van der Waal's radius of Cl atom

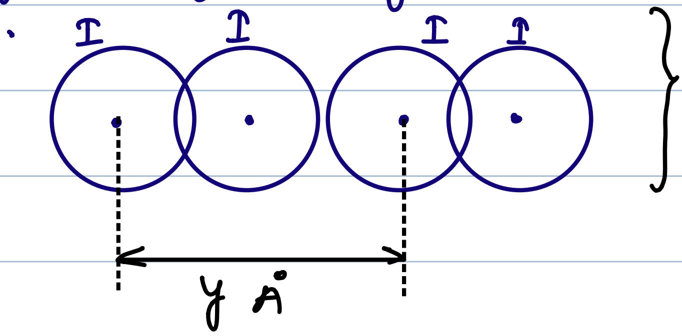
✓ **For noble gases,**

Solid He:



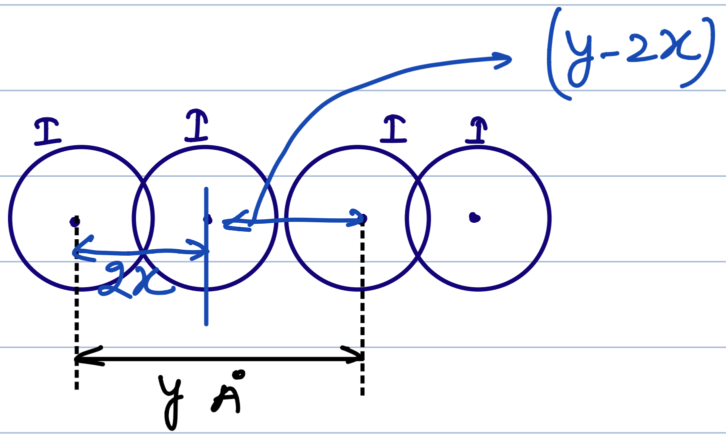
$$\boxed{(r_v)_{He} = \frac{d}{2}}$$

Q. covalent radius of iodine is x Å. find van der waal's radius of iodine considering the following diagram:



closest molecules of solid iodine

solution:



$$(r_v) = \frac{(y-2x)}{2}$$

⁂

If a particular element can show all these three types of radius, then: $r_v > r_m > r_c$

Factors affecting atomic radius:

(i). Number of shells: Generally, greater the no. of shells, more will be the atomic radius.

ex: Li : $1s^2 2s^1$: 2 shells

Na : $1s^2 2s^2 2p^6 3s^1$: 3 shells

Size of Na is more than that of Li.

(ii). Effective nuclear charge:

Generally, more Zeffective → smaller size

ex:

B : $1s^2 2s^2 2p^1$ | C : $1s^2 2s^2 2p^2$ | N : $1s^2 2s^2 2p^3$

$\longrightarrow$

✓ Inc. Zeffective

✓ dec. size

(iii). Bond multiplicity:

For the same bond, Greater the bond multiplicity, smaller the atomic size.

ex: C−C   C=C   C≡C

$\longrightarrow$

✓ inc. bond multiplicity

✓ dec. size of C

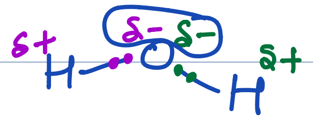
(iv) oxidation number :

Real or hypothetical charge present on atom in its compounds is oxidation no.

$\overset{+}{Na}\overset{-1}{Cl}$ → ox. no. of Na = +1

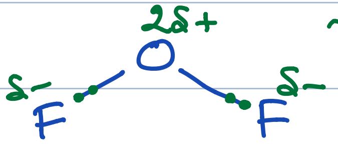
→ ox. no. of Cl = −1

Electronegativity order : $F > O > N \simeq Cl > \underset{\underset{H}{\vee}}{C}$



ox. no. of O = −2

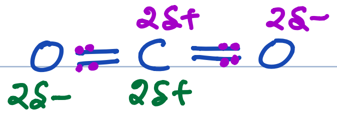
ox. no. of each H = +1



ox. no. of O = +2

ox. no. of each F = −1

CO_2 :



ox. no. of C = +4

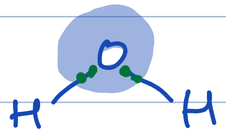
ox. no. of each O = −2

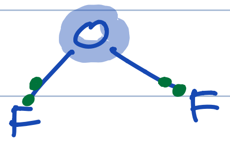
✓ For the same element,

more +ve ox. no. ⟶ smaller size

more −ve ox. no. ⟶ larger size

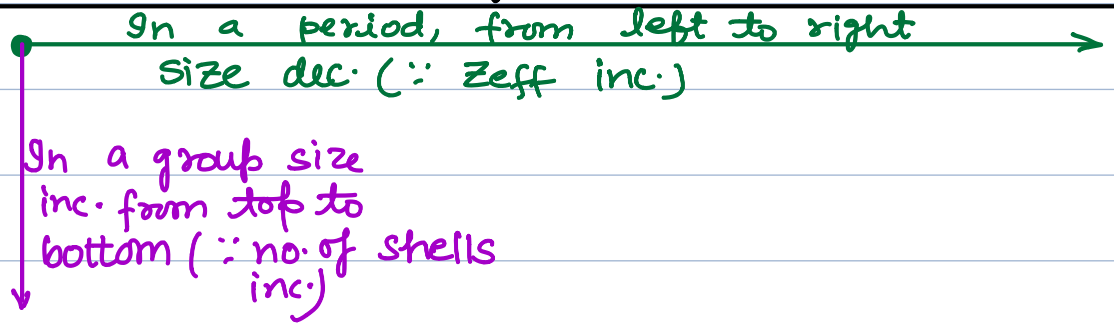
$H_2\overset{-2}{O}$, $\overset{+2}{O}F_2$





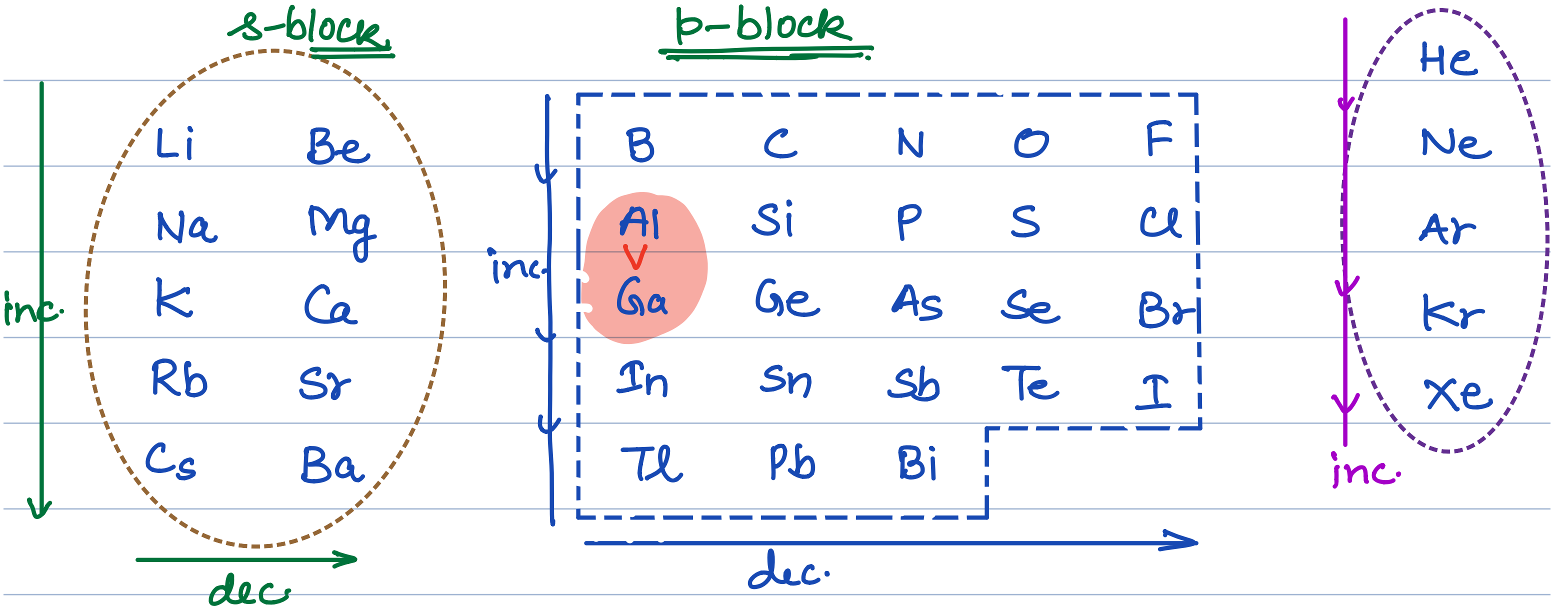
size of oxygen of water is larger than oxygen of OF_2.

General variation of atomic size in periodic table:



In a period, from left to right size dec. (∵ Zeff inc.)

In a group size inc. from top to bottom (∵ no. of shells inc.)



Size of Al and Ga : (v.v. imp).

Atomic radius of Al = 1.43 Å
Atomic radius of Ga = 1.35 Å
} Don't try to learn values.

✓ The size of Ga is slightly smaller than that of Al (If we show leniency in comparison, we can say that size of Ga is nearly equal to that of Al).

Al (13) : [Ne] $3s^2 3p^1$

Ga(31) : [Ar] $4s^2 \underbrace{3d^{10}} 4p^1$

provides poor shielding to outermost e^-

⇓

size dec. ← inc. Z_{eff}

"Transition contraction"

d-block:

size dec. → size almost const → inc →

Gr 3	Gr 4	Gr 5	Gr 6	Gr 7	Gr 8	Gr 9	Gr 10	Gr 11	Gr 12
Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn
∧	∧	∧	∧	∧	∧	∧	∧	∧	∧
Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd
∧	≃	≃	≃	≃	≃	≃	≃	≃	≃
La (f^0)	Hf (72) $[\overset{54}{Xe}]6s^2$ $4f^{14}5d^2$	Ta	W	Re	Os	Ir	Pt	Au	Hg

explanation of variation in period:

Sc —— Cr	inc in Zeffective is dominating
Cr —— Cu	inc. in Z is compensated by inc. in $e^- - e^-$ repulsion
Cu —— Zn	inc. in $e^- - e^-$ repulsion is dominating

explanation of variation of size in groups:

Sc – Zn : 3d-series
Y – Cd : 4d-series
La – Hg : 5d-series

Lanthanoid contraction ⇒

From Gr-4 to Gr : 12

5d-series element ⇒ $\underbrace{4f^{14}\text{ electrons}}_{\text{very poor shielding}}$

very poor shielding → inc. Zeff. → size contraction within a group, size of 5d-series element is nearly equal to that 4d-series element from Gr-4 to Gr-12.

(II). Ionic radius:

[A] Cation radius:

Cations are formed by loss of one or more electrons in atoms.
(i.e. oxidation of atoms)

	$Li \longrightarrow$	$Li^+ + e^-$
No. of $\bar{e}$s:	3	2
No. of protons: (atomic no.)	3	3

In Li & Li^+, Z is same & σ - different

Zeffective is more for Li^+ than for Li.

size of a cation is always smaller than its parent atom because cation has more Zeffective as compared to parent atom.

[B] Anion radius:

Anions are formed by gain of one or more electrons in atom. (i.e. Reduction of atom)

	$F + e^- \longrightarrow$	F^-
No. of $\bar{e}$s :	9	10
No. of protons : (atomic no Z)	9	9

✓ Zeffective more for F atom than for F^- ion.

size of an anion is always larger than its parent atom because anion has lesser Zeffective as compared to parent atom.

✓ comparisons of size:

(i) $O^{2-} > O^{-} > O > O^{+} > O^{2+}$

(ii) $I^{-} > I > I^{+}$

(iii) $Fe > Fe^{+2} > Fe^{+3}$

(iv) $H^{-} > H > H^{+}$ → (proton)

comparison of size in isoelectronic species:

(i):

	O^{2-}	F^{-}	Na^{+}	Mg^{2+}
No. of $\bar{e}$s :	10	10	10	10
No. of protons :	8	9	11	12

Zeffective: $Mg^{2+} > Na^{+} > F^{-} > O^{2-}$

Size: $O^{2-} > F^{-} > Na^{+} > Mg^{2+}$

(ii).

	K^{+}	S^{2-}	Cl^{-}
no. of e^{-}s :	18	18	18
no. of protons :	19	16	17

Zeffective: $K^{+} > Cl^{-} > S^{2-}$

Size: $S^{2-} > Cl^{-} > K^{+}$

v.imp:

"For isoelectronic ions, Greater the positive charge, smaller will be size and Greater the negative charge, larger will be size."

ionic radius comparisons:

- Li^+, Na^+, K^+, Rb^+, Cs^+ $\downarrow$ size inc.
- Be^{2+}, Mg^{2+}, Ca^{2+}, Sr^{2+}, Ba^{+2} $\downarrow$ size inc.
- Ga^+, In^+, Tl^+ $\downarrow$ inc. size
- F^-, Cl^-, Br^-, I^- $\downarrow$ size inc.

In same group, for same charge, size inc. down the group.

- $Li^+ < Be^{2+}$ (Zeff.)
  $Li^+ > Be^{2+}$ (size)
- $N^{3-} < O^{2-}$ (Zeff.)
  $N^{3-} > O^{2-}$ (size)

v. v. imp. trends of ionic radius to remember:

(i)* $\underbrace{I^-}_{5\text{ shells}} > \underbrace{H^-}_{1\text{ shell}} > \underbrace{Br^-}_{4\text{ shells}} > \underbrace{Cl^-}_{3\text{ shells}} > \underbrace{F^-}_{2\text{ shells}}$ } H^- is very large due to low Zeffective.

$H^- \begin{cases} e^- = 2 \\ p = 1 \end{cases}$

(ii)*** $\underbrace{Li^+}_{1\text{ shell}} > \underbrace{Mg^{2+}}_{2\text{ shells}} > \underbrace{Al^{3+}}_{2\text{ shells}}$ } Li^+ is very large due to low Zeffective

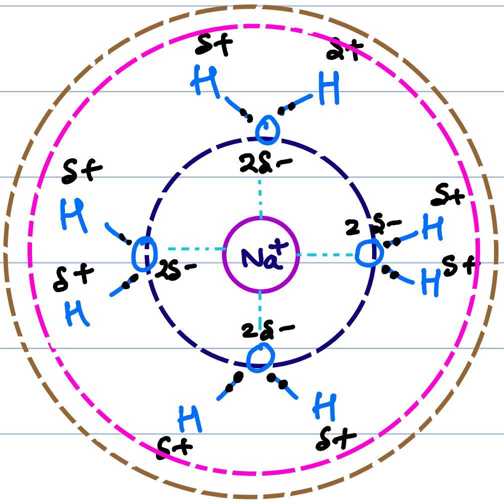
Hydrated radius of ions:

* * * * *

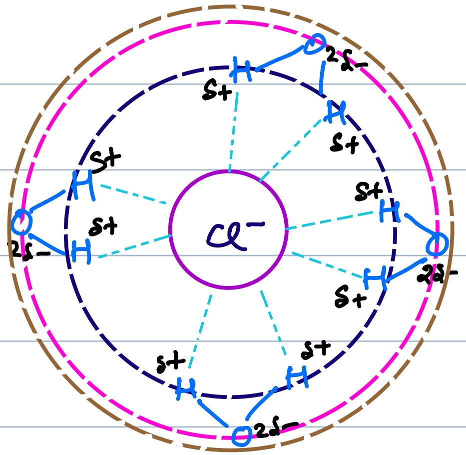
- Hydration describes the interaction of water molecules with the dissolved species.

- when an ionic compound is dissolved into water, it breaks into ions and ions interact with the water molecules.

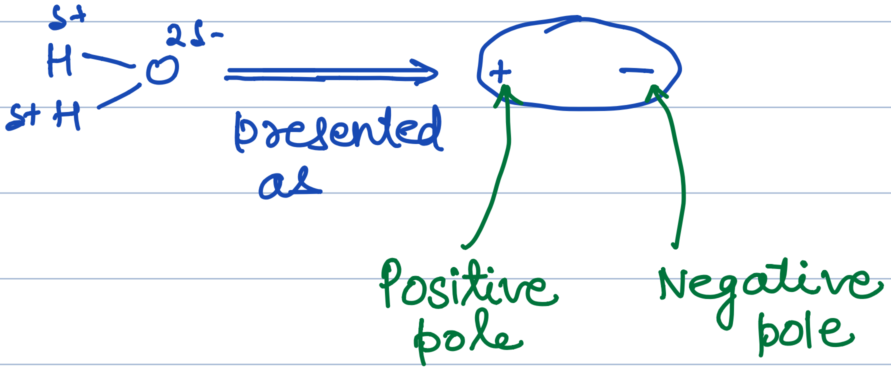
$$NaCl\ (s) \xrightarrow{H_2O} Na^+_{(aq.)} + Cl^-_{(aq.)}$$



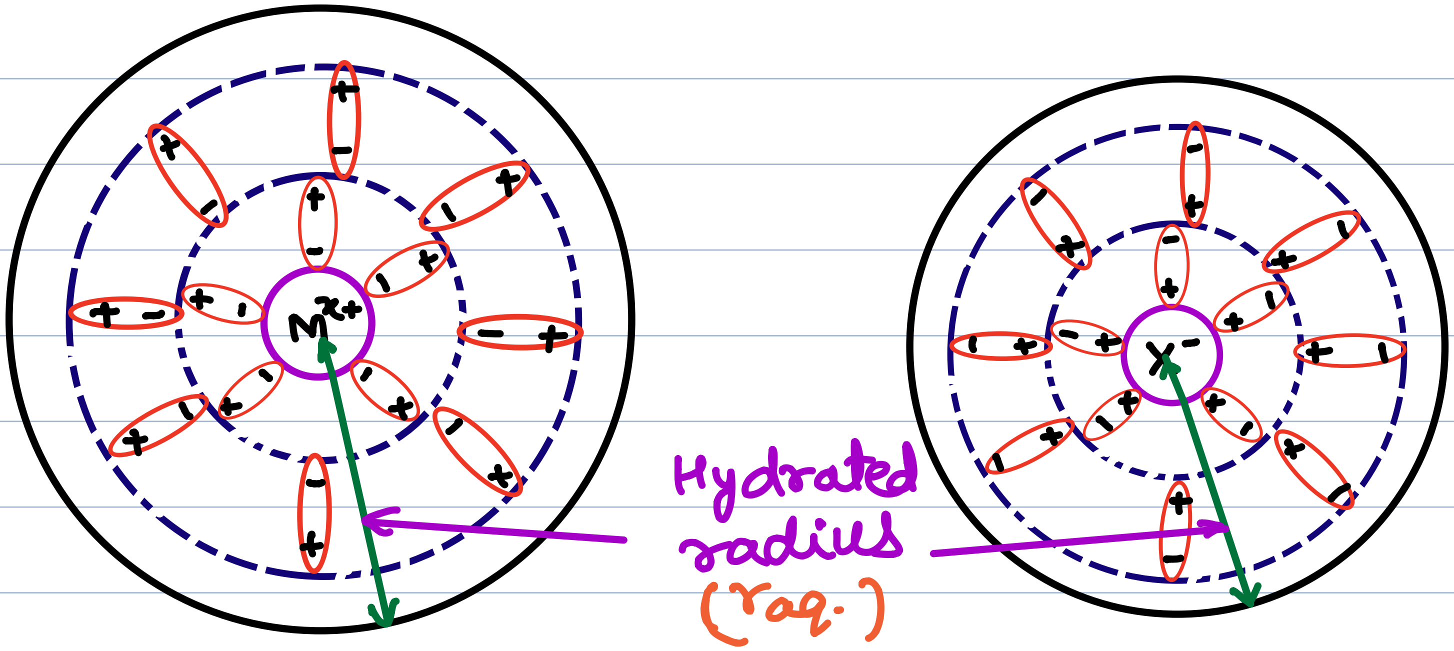
$Na^+(aq.)$
(Hydrated cation)



$Cl^-(aq.)$
(Hydrated anion)



- The energy released on hydration of species is called Hydration energy.



Distance from nucleus of an ion To the point upto which the associated water molecules are immobile due to attraction from ion is called Hydrated radius of that ion.

✓ Extent of hydration indicates the no. of layers of H_2O molecules attatched to a particular ion.

- Greater the charge density of ion, more will be the extent of hydration for the ion.

  ✓ charge density $= \left(\frac{\text{charge}}{\text{volume}}\right)$

- More extent of hydration leads to larger hydrated radius.

ex :

	$Li^+(g)$ $\quad$ $Na^+(g)$ $\quad$ $K^+(g)$ $\quad$ $Rb^+(g)$
size :	$\xrightarrow{\text{inc.}}$
charge density :	$\xrightarrow[\text{dec.}]{}$
extent of hydration :	$\xrightarrow[\text{dec.}]{}$
hydrated radius :	$\xleftarrow[\text{inc.}]{}$
ionic mobility (speed) in water :	$\xrightarrow[\text{inc.}]{}$
electrical conductance :	$\xrightarrow[\text{inc.}]{}$

$Li^+_{(aq.)} > Na^+_{(aq.)} > K^+_{(aq.)} > Rb^+_{(aq.)}$: size

ex :

	Li^+		Be^{2+}
charge density :		$<$	
extent of hydration :		$<$	
Hydrated radius :		$<$	

ex : $F^-_{(aq.)} > Cl^-_{(aq.)} > Br^-_{(aq.)}$ (size)

(III). Ionisation Energy:

"Minimum amount of energy required to remove the outermost electron from ground state of an isolated gaseous atom is called ionisation energy of atom."

- Ionisation enthalpy and ionisation potential have exactly same trends as that of ionisation energy.

Reactions releasing energy = exothermic reactions

↳ '-'ve $\underbrace{\text{enthalpy change}}_{\Delta H}$

Reaction absorbing energy = endothermic reactions

↳ '+'ve $\underbrace{\text{enthalpy change}}_{\Delta H}$

- For an element,

$$|\text{ionisation energy}| \simeq |\text{ionisation enthalpy}|$$

* Ionisation enthalpy is positive for a neutral atom (i.e. endothermic).

✓ * successive ionisation energy values of an element increases continuously (due to dec. size and inc. Zeff) → $z > y > x$

✓ $$\boxed{Q \xrightarrow{\Delta H = +x} Q^{+} \xrightarrow{\Delta H = +y} Q^{2+} \xrightarrow{\Delta H = +z} Q^{3+}}$$

First ionisation energy of Q (IE_1 of Q) = x

First ionisation energy of Q^+ (IE_1 of Q^+) = y
[Second ionisation energy of Q (IE_2)] = y

First ionisation energy of Q^{2+} (IE_1 of Q^{2+}) = z
[Second ionisation energy of Q^+ (IE_2 of Q^+)] = z
[Third ionisation energy of Q (IE_3 of Q)] = z

- Energy required to remove first two electrons from Q = $(x+y)$

- Energy required to remove first three electrons from Q = $(x+y+z)$

- Energy req. to remove first two e⁻s from $Q^+ = (y+z)$

Factors affecting ionisation energy:

(i) **size**: Generally, smaller the size more will be the ionisation energy.

ex: $Li \underset{size}{>} Be$, $\underset{IE_1}{<}$

(ii) **Zeffective**: Generally, more Zeffective leads to more ionisation energy.

ex $C \underset{Z_{eff}}{<} O$, $\underset{IE_1}{<}$

(iii) **stability of electronic configuration**:

✓ Exactly half-filled and fully filled subshells provide stability to the configuration.

✓ stable configuration leads to more IE.

[This factor is usually dominating upto adjacent elements]

ex: $N : 1s^2\ 2s^2\ 2p^3$ ⁞ $O : 1s^2\ 2s^2\ 2p^4$ ⁞ $F : 1s^2 2s^2 2p^5$


exactly half filled

IE₁ order: $F > N > O$

$$\boxed{\begin{array}{ll} \text{size}: & N > O > F \\ Z_{eff}: & N < O < F \end{array}}$$

(iv) Penetration effect :

Removal of electron from orbital having more penetrating power is more difficult (i.e. more IE)

[This factor is usually dominating upto adjacent elements]

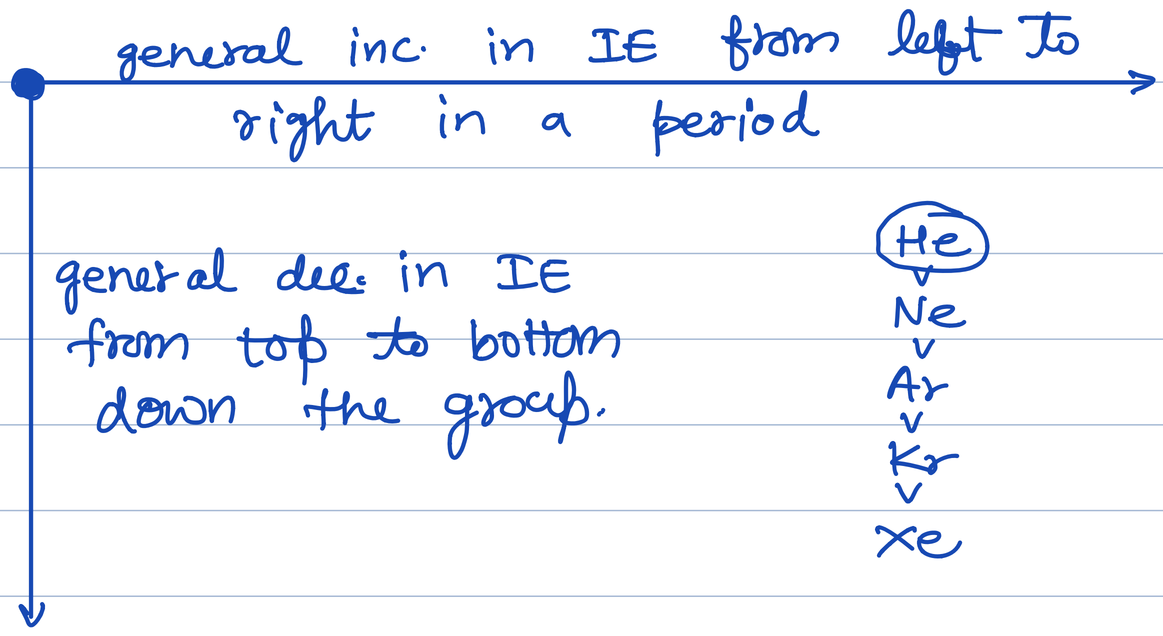
ex:

Be $\underset{IE_1}{>}$ B

$1s^2\ 2s^2$ $\boxed{\uparrow\downarrow}$ $\qquad$ $1s^2\ 2s^2\ 2p^1$ $\boxed{\uparrow\,|\;\,|\;\,}$

($\because$ 2s is more penetrating than 2p)

General variation of IE in periodic table :



general inc. in IE from left to right in a period

general dec. in IE from top to bottom down the group.

He
∨
Ne
∨
Ar
∨
Kr
∨
Xe

Periodwise variation :

(i). H $\underset{IE}{<}$ He

$1s^1$ $\qquad$ $1s^2$ (stable configuration) (ii). Li < Be > B < C < N > O < F < Ne

$2s^1$ $\quad 2s^2 \quad 2p^1 \quad 2p^2 \quad 2p^3 \quad 2p^4 \quad 2p^5 \quad 2p^6$

$$\text{Li} < \underbrace{\text{Be} > \text{B}}_{\text{Penetration effect}} < \text{C} < \underbrace{\text{N} > \text{O}}_{\text{stability of configuration}} < \underbrace{\text{F} < \text{Ne}}_{\text{stability of configuration}}$$

Li < B < Be < C < O < N < F < Ne

(iii). Na < Mg > Al < Si < P > S < Cl < Ar

$3s^1$ $\quad 3s^2 \quad 3p^1 \quad 3p^2 \quad 3p^3 \quad 3p^4 \quad 3p^5 \quad 3p^6$

$$\text{Na} < \underbrace{\text{Mg} > \text{Al}}_{\text{Penetration effect}} < \text{Si} < \underbrace{\text{P} > \text{S}}_{\text{stability of configuration}} < \underbrace{\text{Cl} < \text{Ar}}_{\text{stability of configuration}}$$

Na < Al < Mg < Si < S < P < Cl < Ar

IE trends in Gr: 13 and Group: 14 :

बो तल गा लि(यां) इंडिया(में)

$$\boxed{\text{B} > \underbrace{\text{Tl}}_{} > \text{Ga} > \text{Al} > \text{In}}$$

due to high Zeff. (presence of 4f e's)

$$\boxed{\text{C} > \text{Si} > \text{Ge} > \underbrace{\text{Pb}}_{} > \text{Sn}}$$

due to more Zeff (presence of 4f e's)

v. imp. table:

Maximum IE in a period is for	Noble gas element
Least IE in a period is for	Group-1 element
max IE among all Know elements	Helium
Least IE among all Know elements	Cs

d- block :

Gr:3

Sc

↓ dec.

Y

↓

La

Gr:4 to Gr:12

3d- series

4d- series

^

5d- series

∵ size nearly same & more Zeff in 5d

ex { Cd ^ Hg }

Ionisation energy in 3d-series :

Sc ㉑	Ti ㉒	V ㉓	Cr ㉔	Mn ㉕	Fe ㉖	Co ㉗	Ni ㉘	Cu ㉙	Zn ㉚
$4s^2$ $3d^1$	$4s^2$ $3d^2$	$4s^2$ $3d^3$	$4s^1$ $3d^5$	$4s^2$ $3d^5$	$4s^2$ $3d^6$	$4s^2$ $3d^7$	$4s^2$ $3d^8$	$4s^1$ $3d^{10}$	$4s^2$ $3d^{10}$
$4s^0 3d^0$	$4s^0 3d^1$	$4s^0 3d^2$	$4s^0 3d^3$	$4s^0 3d^4$	$4s^0 3d^5$	$4s^0 3d^6$	$4s^0 3d^7$	$4s^0 3d^8$	$4s^0 3d^9$

Maximum IE_1 in 3d-series	Zn (stable configuration)
Maximum IE_2 in 3d-series	Cu (e^- to be removed from $3d^{10}$)
Maximum IE_3 in 3d-series	Zn (e^- to be removed from $3d^{10}$)
Maximum IE_4 in 3d-series	Sc (e^- removal from noble gas config.

Sc^{3+} : $[Ar]\ 4s^0\ 3d^0$

IE_2 and IE_3 comparisons :

- **Compare second ionisation energy in following pairs :**

(a) He, Li (2) (3) (b) Be, B (4) (5) (c) N, O (7) (8) (d) F, Cl (9) (17)

Sol.:

(a) He : $1s^2$ | He^+ : $1s^1$ | → IE_2 : $\boxed{Li > He}$

Li : $1s^2 2s^1$ | Li^+ : $1s^2$ (More stable) | IE_1 : He > Li

(b) Be : $1s^2\ 2s^2$ | Be^+ : $1s^2 2s^1$ | → IE_2 $\boxed{B > Be}$

B : $1s^2\ 2s^2\ 2p^1$ | B^+ : $1s^2\ 2s^2$ (More stable) | IE_1 = $\boxed{Be > B}$

(c) N : $1s^2\ 2s^2\ 2p^3$ | $N^+ : 1s^2\ 2s^2\ 2p^2$ | $\rightarrow$ $\boxed{IE_2 : O > N}$

O : $1s^2\ 2s^2\ 2p^4$ | $O^+ : 1s^2\ 2s^2\ 2p^3$ |

$IE_1 : N > O$

(d)

F : $1s^2\ 2s^2\ 2p^5$ | $F^+ : 1s^2\ 2s^2\ 2p^4$ | $\rightarrow$ $IE_2 : F > Cl$

Cl : $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$ | $Cl^+ : 1s^2 2s^2 2p^6 3s^2\ 3p^4$ | $IE_1 : F > Cl$

Q Compare IE_2 & IE_1 for C and N

Sol:

C(6) : $1s^2\ 2s^2\ 2p^2$

N(7) : $1s^2\ 2s^2\ 2p^3$

$IE_1 : \boxed{N > C}$

$C^+ : 1s^2\ 2s^2\ 2p^1$

$N^+ : 1s^2\ 2s^2\ 2p^2$

$IE_2 : \boxed{N > C}$

- **Compare third ionisation energy in following pairs:**

✓(a) Li (3) , Be (4) ✓(b) B, (5) C (6) ✓(c) O, (8) F (9)

Sol:

(a) $Li^{2+} : 1s^1$
$Be^{2+} : 1s^2$
IE_3 : Be > Li

(b) $B^{2+} : 1s^2 2s^1$
$C^{2+} : 1s^2 2s^2$
IE_3 : C > B

(c) $O^{2+} : 1s^2 2s^2 2p^2$
$F^{2+} : 1s^2 2s^2 2p^3$
IE_3 : F > O

** Ionisation energy comparisons between atoms and ions of same element :

******** [Size & Zeffective are dominating over configuration]

(i) $\underset{1s^1}{H} \underset{IE_1}{>} \underset{1s^2}{H^-}$

(ii) $\underset{IE_1}{\underline{Cl^+ > Cl > Cl^-}}$

$\boxed{\text{IE of Cation > IE of atom > IE of anion}}$ → for the same element.

Applications of ionisation energy :

[1]. Metallic character :

Lower the IE ⟶ More electropositivity ⟶ more metallic character

Lower the IE ⟶ Generally more reactive metal

ex :

Li
Na
K
Rb
Cs

↓ dec. IE
↓ inc. metallic character
↓ inc. reactivity

$Li \underset{IE}{<} Be$

$Li \underset{reactivity}{>} Be$

* $\boxed{\text{Alkali metal} \underset{IE}{<} \text{d-block metal}}$

[2] Prediction of group no. of s and p-block elements on basis of successive IE data:

***
**

Maximum difference in	no. of valence e^-s	Group No.
IE_1 and IE_2	1	1
IE_2 and IE_3	2	2
IE_3 and IE_4	3	13
IE_4 and IE_5	4	14
IE_5 and IE_6	5	15
IE_6 and IE_7	6	16
IE_7 and IE_8	7	17

ex (i) element 'x': IE_1 IE_2 IE_3

5.1 eV/atom 47 eV/atom 69 eV/atom

diff = 47 - 5.1 ≃ 42 eV diff = 22 eV

Biggest Jump in IE is from IE_1 to IE_2.

No. of valence e^- = 1

Group no = 1

ex (ii) element 'Z':

IE_1 IE_2 IE_3 IE_4 IE_5

8 eV/atom 24 eV/atom 60 eV/atom 250 eV/atom 352 eV/atom

Sol

diff = 16 eV diff = 36 eV dif 190 eV diff = 102

No. of valence e^- = 3

Group no. = 13

Jee-advanced-2020:

The 1st, 2nd, and the 3rd ionization enthalpies, I_1, I_2, and I_3, of four atoms with atomic numbers n, $n+1$, $n+2$, and $n+3$, where $n < 10$, are tabulated below. What is the value of n?

Atomic number	Ionization Enthalpy (kJ/mol)		
	I_1	I_2	I_3
n	1681	3374	6050
$n+1$	2081	3952	6122
$n+2$	496	4562	6910
$n+3$	738	1451	7733

✓ Alkali metal → $n+2$

Sol: → at. no. less than 12

Li $(n=1)$ at. no. $=3$

Na $(n=9)$ at. no. $=11$

ANS: 9

when IE data are given and noble gas is asked, → The one having maximum IE_1 among given will be assigned as noble gas.

[3]. Stability of oxidation states:
*

successive IE difference ≥ 16 eV	successive IE difference ≤ 11 eV	successive IE diff. between 11 to 16 eV
⇓ Lower ox. state is more stable	⇓ Higher ox. state is more stable	⇓ Both ox. states are stable
ex: IE_1 of Na = 5.1 eV IE_2 of Na = 47 eV $\underset{\text{(stable)}}{Na^+} > Na^{2+}$	**ex:** IE_2 of Al = 18 eV IE_3 of Al = 28 eV $\underset{\text{(stable)}}{Al^{3+}} > Al^{2+}$	**ex.** IE_2 of Fe = 16 eV IE_3 of Fe = 30 eV Both Fe^{2+} and Fe^{3+} are stable.

[4]. Reducing power:

Generally lower the IE, more is reducing power (provide e^- for reduction of others).

ex

Be
Mg ↓ dec. IE
Ca ↓ inc. reducing power
Sr ↓
Ba ↓

$Cs \underset{IE_1}{<} Na$

$\underset{\text{Reducing power.}}{>}$

[D]. Electron affinity, Electron gain enthalpy:

Electron affinity (EA)	Electron gain enthalpy ($\Delta_{eg}H$)
The amount of energy released on addition of an electron to the outer most shell of isolated gaseous atom is electron affinity of that atom.	The enthalpy change on addition of an electron to outermost shell of isolated gaseous atom is electron gain enthalpy of that atom.

ex:

$$A(g) + e^- \longrightarrow A^-(g) \; ; \; \Delta H = -x \text{ eV}$$

$$\left\{ \begin{array}{l} \Delta_{eg}H \text{ for } A = -x \\ EA \text{ for } A \simeq x \end{array} \right.$$

$$\boxed{\Delta_{eg}H \simeq -EA}$$

ex:

$$Z_{(g)} + e^- \longrightarrow Z^-(g) \, , \; \Delta H = +y \text{ eV}$$

$$\left\{ \begin{array}{l} \Delta_{eg}H \text{ for } Z = +y \\ EA \text{ for } Z \simeq -y \end{array} \right.$$

ex:

$$\begin{cases} A(g) + e^- \longrightarrow A^-(g); & \Delta H = -10, \; EA \simeq 10 \\ D(g) + e^- \longrightarrow D^-(g); & \Delta H = -15, \; EA \simeq 15 \end{cases}$$

Explanation:

✓ Electron affinity of D is more than that of A.

✓ D has more -ve electron gain enthalpy than that of A.

$$X(g) \xrightarrow[\Delta H_1]{+e^-} X^{\ominus}(g) \xrightarrow[\Delta H_2]{+e^-} X^{2-}(g)$$

$\Delta H_1 \rightarrow \Delta_{eg_1}H$ of X

$\Delta H_2 \rightarrow \Delta_{eg_1}H$ of $X^{\ominus}$

$\Delta H_2 \rightarrow \Delta_{eg_2}H$ of X

very important points:

*** [i]. $\Delta_{eg_1}H$ for s and p-block elements

positive for	Negative for
✓ Noble gas elements (He, Ne, Ar, Kr, Xe, ...) $ns^2 np^6$	✓ Rest elements of s and p-block
✓ Group-2 metals (Be, Mg, Ca, Sr, Ba--) ns^2	
✓ Nitrogen element $2p^3$ (+ more interelectronic repulsion]	

**
[ii]. Addition of an electron to a cation is "Exothermic" ∵[attraction b/w e^- being added & ion]

$$\boxed{\Delta_{eg_1}H \text{ of a Cation} = \text{'-'ve}}$$

**
[iii]. Addition of an electron to an anion is "Endothermic" ∵[Repulsion b/w e^- being added & ion]

$$\boxed{\Delta_{eg_1}H \text{ of an anion} = \text{'+'ve}}$$

(∵ $\Delta_{eg_2}H$, $\Delta_{eg_3}H$, ---- are positive for any atom)

****
[iv]. Addition of two or more electrons to an atom is always endothermic.

ex :

$$X(g) + e^- \longrightarrow X^{\ominus}(g) \quad , \Delta H_1 = -p_1$$
$$X^{\ominus}(g) + e^- \longrightarrow X^{2-}(g), \quad \Delta H_2 = +p_2$$

$$X(g) + 2e^- \longrightarrow X^{2-}(g), \quad \Delta H_T = (p_2 - p_1)$$

↓

+ve (∵ $|p_2| > |p_1|$)

ex -

$$O(g) + 2e^- \longrightarrow O^{2-}(g), \quad \Delta H = \text{'+'ve}$$
$$N(g) + 3e^- \longrightarrow N^{3-}(g), \quad \Delta H = \text{'+'ve}$$

Fill the following table:

Species	Process	Sign of ΔH
Atom	Removal of 1 or more electrons	
Atom	Addition of 1 electron	
Atom	Addition of 2 or more electrons	
Cation	Removal of 1 or more electrons	
Cation	Addition of 1 electron	
Cation	Addition of 2 or more electrons	
Anion	Addition of 1 or more electrons	
Anion	Removal of 1 or more electrons	

Factors affecting electron affinity:

(i). **Size:** smaller the size of atom, more attractive force on coming electron, generally more energy is released.

(ii). Effective nuclear charge:

more Zeffective, generally more energy released.

(iii). Stability of configuration:

exactly half-filled or fully filled configuration ⟶ less tendency to accept e^- [either +ve $\Delta_{eg}H$ or $\underbrace{-ve\ \Delta_{eg}H}_{\text{Less magnitude}}$

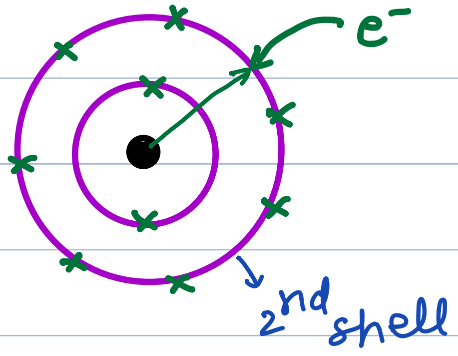
**

(iv). Interelectronic repulsion:

More interelectronic repulsion may lead to lower electron affinity.

ex :

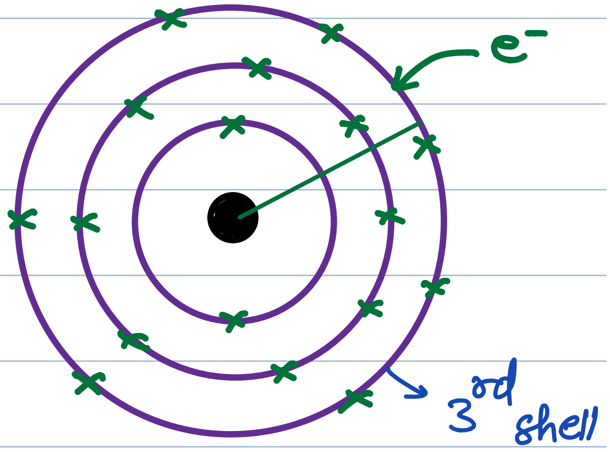
F(9) : 2,7



✓ more e^- density in the shell to which electron is being added → more repulsion to coming electrons by already present electrons → less energy release

[Less EA]

Cl(17) : 2, 8, 7



Less e^- density in the shell to which e^- is being added,
↳ less repulsion to coming e^-.
↓
[More EA]

EA in p-block :

✓ 2nd Period	B	C	N	O	F
	∧	∧	∧	∧	∧
✓ 3rd Period	Al	Si	P	S	Cl

Due to more interelectronic repulsions, the second period elements have lower EA than corresponding third period elements in p-block.

Electron affinity orders to remember :

(i) Group No: 17 : $Cl > F > Br > I$

(ii) Group No: 16 : $S > Se > Te > Po > O$

EA comparison in Periods:

(i) $H > He$

(ii) $Li > \underset{2s^2}{Be} < B < C > \underset{2p^3}{N} < O < F > \underset{2p^6}{Ne}$

(iii) $Na > \underset{3s^2}{Mg} < Al < Si > \underset{3p^3}{P} < S < Cl > \underset{3p^6}{Ar}$

- In respective periods, generally halogens have maximum electron affinity. (Reason: high Zeff smaller size)

- "Cl has maximum electron affinity among all known elements."

- For same period,

$$\left[\underset{ns^1}{\text{Group-1 metal}} \underset{EA}{>} \underset{np^1}{\text{Group-13 element}}\right]$$

ex: $\left\{\begin{array}{l} Li \underset{EA}{>} B \\ Na \underset{EA}{>} Al \end{array}\right.$

Some v. imp points :

(i) $\boxed{|\Delta_{eg}H \text{ of atom}| = |\Delta_{\text{ionisation}}H \text{ of anion}|}$

→ For same element & its anion.

ex :

$$\begin{cases} Cl + e^- \longrightarrow Cl^-; & \Delta H = -x \\ Cl^- \longrightarrow Cl + e^-; & \Delta H = +x \end{cases}$$

; [e⁻ gain of Cl atom

; [ionisation of Cl^- ion

(ii) $\boxed{|\Delta_{\text{ionisation}}H \text{ of atom}| = |\Delta_{eg}H \text{ of cation}|}$

→ For same element & its cation

ex :

$Na \longrightarrow Na^+ + e^-$, $\Delta H = +y$ [ionisation of Na atom

$Na^+ + e^- \longrightarrow Na$, $\Delta H = -y$ [e⁻ gain of Na^+ ion

****

(iii) $\boxed{IE_1 \text{ of } F^- < IE_1 \text{ of } Cl^-}$

Similarly, [IE_1 of $O^- <$ IE_1 of S^-]

explanation:

$F + e^- \longrightarrow F^-$, $\Delta H = -x$

$Cl + e^- \longrightarrow Cl^-$, $\Delta H = -y$

$|y| > |x|$

$F^- \longrightarrow F + e^-$, $\Delta H = +x$

$Cl^- \longrightarrow Cl + e^-$, $\Delta H = +y$

IE_1 of $Cl^- >$ IE_1 of F^-

IE_1 : $F > Cl > Cl^- > F^-$

[while comparing IE_1 of anions from same group, use reverse of Electron affinity]

- while comparing EA of cations of same group, use Reverse of IE.

ex:

$F \longrightarrow F^+ + e^-$, (more absorb) } $F^+ + e^- \longrightarrow F$ (more release)

$Cl \longrightarrow Cl^+ + e^-$, } $Cl^+ + e^- \longrightarrow Cl$

$\left[F^+ \underset{EA}{>} Cl^+ \right]$

Species	Process	Sign of ΔH
Atom	Removal of 1 or more electrons	+'ve
Atom	Addition of 1 electron	+'ve or -'ve
Atom	Addition of 2 or more electrons	+'ve
Cation	Removal of 1 or more electrons	+'ve
Cation	Addition of 1 electron	-'ve
Cation	Addition of 2 or more electrons	+'ve or -'ve
Anion	Addition of 1 or more electrons	+'ve
Anion	Removal of 1 or more electrons	+'ve or -'ve

Q. $O \longrightarrow O^{2-}$ or $N \longrightarrow N^{3-}$ is endothermic. Still oxides (O^{2-}) or nitrides (N^{3-}) are abundant in nature. Why?

Sol: $O \xrightarrow[\text{endo}]{} O^{2-}$

compensates for energy given via energy released on bond formation.

$N \xrightarrow[\text{endo}]{} N^{-3}$

compensates for energy given via energy released on bond formation.

Jee-advanced

Q which of the following ion is the least stable?

(A) Li^- (B) C^- (C) Be^- (D) F^-

Sol: ANS: (C)

$$\left\{\begin{array}{lll} Li + e^- & \longrightarrow Li^- & (\text{exo}) \\ C + e^- & \longrightarrow C^- & (\text{exo}) \\ F + e^- & \longrightarrow F^- & (\text{exo}) \end{array}\right.$$

$$\left\{ Be + e^- \longrightarrow Be^- \ (\text{endo})\right.$$

Multiple correct

Q. Which of the following processes are exothermic

(A) $O^{2-}(g) \longrightarrow O(g)$ (B) $Be^-(g) \longrightarrow Be(g)$

(C) $P(g) \longrightarrow P^{-3}(g)$ (D) $H^+(g) \longrightarrow H^-(g)$

Sol: [ANS: (A,B,D)]

Ⓐ $O \longrightarrow O^{2-}$ (endo) ; $O^{2-} \longrightarrow O$ (exo)

$Be \longrightarrow Be^-$ (endo) ; $Be^- \longrightarrow Be$ (exo)

$P \longrightarrow P^{-3}$ endo

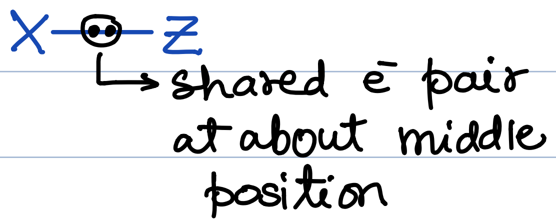
$H^+ + e^- \longrightarrow H$ (exo)

$H + e^- \longrightarrow H^-$ (exo)

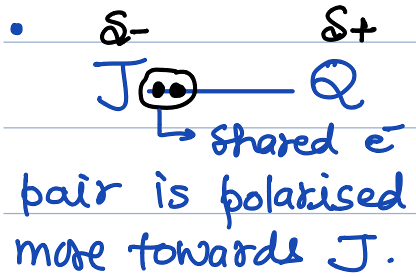
$H^+ + 2e^- \longrightarrow H^-$ (exo)

[E]. Electronegativity (χ) :

The tendency of a covalently bonded atom to attract the shared electron pairs towards itself is electronegativity.

- 

Electronegativity of X is equal electronegativity of Z.

- 

Electronegativity of J is more than that of Q.
[Due to this electronegativity difference, polarity is arised in the bond]

Characteristics of electronegativity :

(i). It is property of bonded atom not of isolated atom.

(ii). It is unitess, dimensionless.

(iii). It is a variable property, relative property.

Scales of electronegativity : (only for Jee-advanced, KVPY, olympiad)

Most followed

(i) Pauling's scale :

✓ based on bond energy data

Let :

$$\overset{\delta+}{A}\text{———}\overset{\delta-}{B}$$

$$\text{Resonance energy}(\Delta) = \text{Experimental bond dissociation energy }(E_{A-B}) - \text{Theoretical bond dissociation energy } \underbrace{\sqrt{E_{A-A} \times E_{B-B}}}_{\text{Geometric mean}}$$

✓

$$\boxed{|\chi_B - \chi_A| = \sqrt{\Delta}}$$

(Δ in ev/atom)

Reference:

Electronegativity of F = 4.0 [acc. to NCERT]

$$\boxed{1 \text{ ev/atom} = 96.46 \text{ KJ/mol}}$$

[1 cal = 4.18 J]

EN values on Pauling's scale:

$H^{2.1}$						
$Li^{1.0}$	1.5 Be	2.0 B	2.5 C	3.0 N	3.5 O	$F^{4.0}$
		Al 1.5	Si 1.8	P 2.5	S 2.8	$Cl^{3.0}$
						$Br^{2.8}$
						$I^{2.5}$

*

$$F > O > N \simeq Cl > C \simeq I > \cdots > Cs$$

(ii) Mulliken's scale:

$$\boxed{\chi_M = \frac{(IE + EA)}{2}}$$

→ IE and EA in eV/atom

For an element:

$$\boxed{\chi_M \simeq 2.8\,\chi_p}$$

χ_M: Mulliken's EN
χ_p: Pauling's EN

$$\boxed{\chi_p = \frac{(IE + EA)}{2 \times 2.8}}$$

→ IE & EA in eV/atom

[iii]. Allred-Rochow scale:

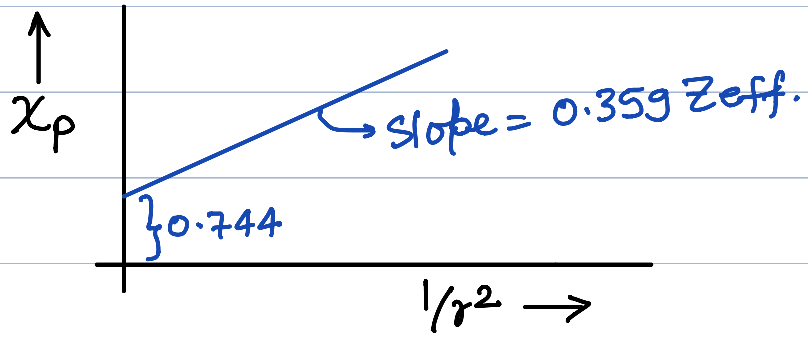
$$\boxed{\chi_{AR} = \frac{0.359\, Z_{eff.}}{r^2}}$$

Z_{eff} = effective nuclear charge
r = covalent radius (Å)

For an element,

$$\boxed{\chi_p \simeq 0.744 + \chi_{AR}}$$

$$\boxed{\chi_p \simeq 0.744 + 0.359\,\frac{Z_{eff}}{r^2}}$$



Factors affecting electronegativity :

(i). **Size :** smaller the size, generally more EN.

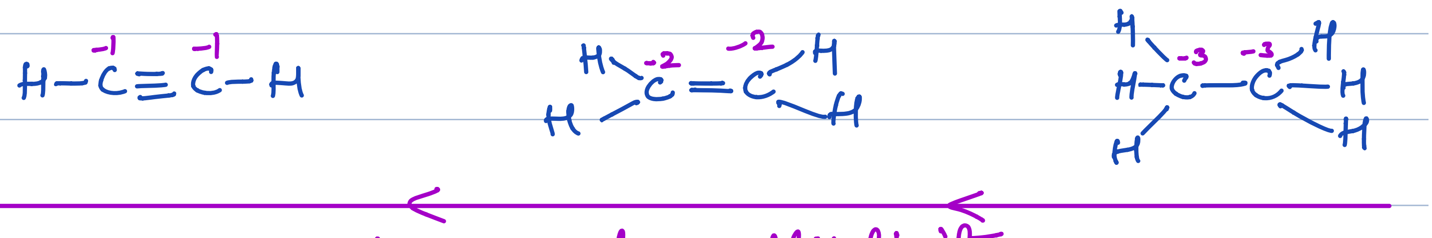
(ii). **Zeffective :** More Zeffective, generally more EN.

(iii). **Oxidation no.** For same element,

{ more +ve ox. no. ⟹ more EN
{ more −ve ox. no. ⟹ less EN

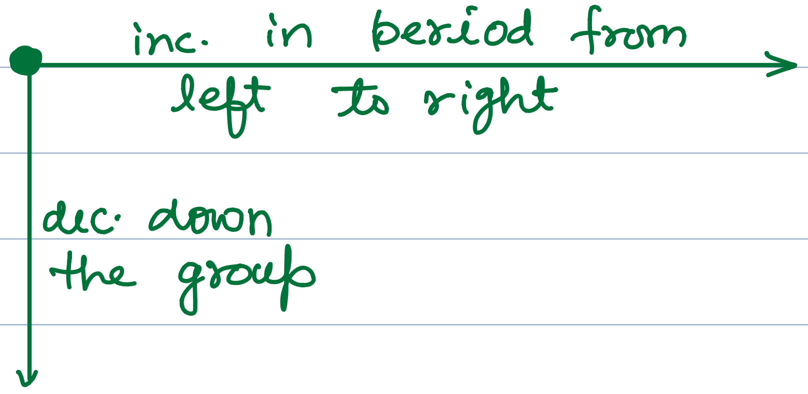
(iv). **Bond multiplicity (% s-character) :** → in next chapter

for same element, more bond multiplicity, more EN.



inc. bond multiplicity
inc. EN of C

General trends in EN :



2^{nd} period:

Li < Be < B < C < N < O < F

3^{rd} period:

Na < Mg < Al < Si < P < S < Cl

In any period:

(≠ Period=1)

most EN → Halogen

Least EN → Alkali metal

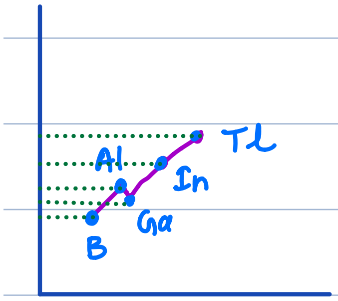
*** EN order to remember:

$$F > O > N \simeq Cl > \underbrace{C \simeq S \simeq I}_{\substack{S > C \\ \downarrow \\ \text{slightly}}} > \underbrace{P \simeq H}_{\substack{P > H \\ \downarrow \\ \text{slightly}}} > \cdots\cdots > Cs$$

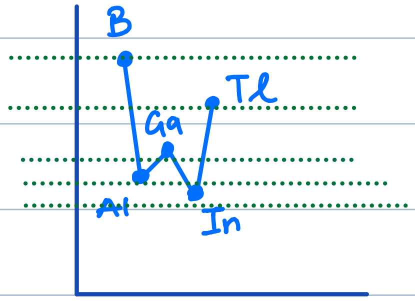
p- Block Groups which do not follow general variation of EN: (not so imp.)

Gr: 13	Gr: 14
B (2.0)	C (2.5)
Al (1.5)	Si (1.8)
	$\simeq$
Ga (1.6)	Ge (1.8)
	$\simeq$
In (1.7)	Sn (1.8)
	$\simeq$
Tl (1.8)	Pb (1.8)

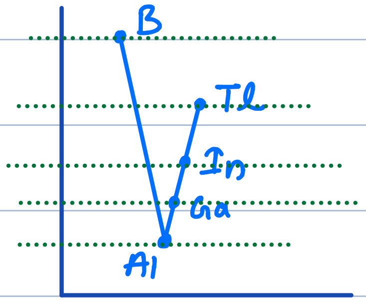
size



IE_1



EN



Applications of EN:

(i) % ionic character in a covalent bond:

$$\overset{\delta+}{A} — \overset{\delta-}{B}$$

more EN difference b/w A and B, more partial charge developed on them ⇒ more ionic character in bond.

ex: $\overset{\delta+}{H}-\overset{\delta-}{F}$ $\quad$ $\overset{\delta+}{H}-\overset{\delta-}{Cl}$ $\quad$ $\overset{\delta+}{H}-\overset{\delta-}{Br}$

←←

inc. EN difference

inc. % ionic character in bond

(ii) Metallic and non-metallic nature :

EN : directly related to non-metallic character

EN : inversely related to metallic character

Q. compare non-metallic nature of

K, Si, Cl

sol: K < Si < Cl (as per EN trend)

Nature of oxides:

[A]. Acidic oxides: oxides which react with bases to form salt are acidic oxides.

- * Most of the non-metal oxides are acidic.

ex: CO_2, SO_2, SO_3, Cl_2O_7, B_2O_3, SiO_2, NO_2, N_2O_5, ...

- oxides of metals or metalloids in which ox. no. of metal or metalloid is greater or equal to +5 are acidic.

ex: $\overset{+6}{Mn}O_3$, Mn_2O_7, $\overset{+6}{Cr}O_3$, Bi_2O_5, As_2O_5, Sb_2O_5 ...

Mn_2O_7 → +7 each; Bi_2O_5 → +5 each; As_2O_5 → +5 each; Sb_2O_5 → +5 each

(ox. no. of each oxygen in oxide = −2)

Comparison of acidic strength in oxides:

[more EN ⟶ more non-metallic nature ⟶ more acidic]

(i) $\overset{+4}{S}O_2 < \overset{+6}{S}O_3$

(ii) $SO_2 > CO_2$
(S slightly more EN than C)

(iii) $\overset{+5}{N_2}O_5 > \overset{+4}{N_2}O_4 > \overset{+3}{N_2}O_3$

(iv) $P_4O_{10} < SO_3 < Cl_2O_7$
(EN: P<S<Cl)

(v) $CO_2 < N_2O_5$
(EN: N>C)

[B]. Basic oxides:

oxides which react with acids to form salt are basic oxides

- Metal oxides in which ox. no. of metal is less than +5 are generally basic.

ex: $\overset{+1}{Na_2O}$, $\overset{+2}{CaO}$, $\overset{+3}{Fe_2O_3}$, $\overset{+2}{CuO}$,
$\overset{+1}{Cu_2O}$, $\overset{+2}{HgO}$, $\overset{+2}{MnO}$, $\overset{+2}{CrO}$
$\overset{+2}{FeO}$

comparison of basic strength in oxides:

[less EN ⟶ more metallic nature ⟶ more basic]

(i) $Li_2O < Na_2O < K_2O < Rb_2O < Cs_2O$

(ii) $Na_2O > MgO$

(iii) $Na_2O > CuO$

(iv) $\overset{+2}{FeO} > \overset{+3}{Fe_2O_3}$

[C]. Amphoteric oxides:

oxides which react with both acids and bases are amphoteric oxides.

List of amphoteric oxides:

s-block

BeO

p-block

Gr:13	Gr 14	Gr 15
Al_2O_3	SnO, SnO_2	
Ga_2O_3	PbO, PbO_2	$\overset{+3}{As}_2O_3$
	Pb_2O_3, Pb_3O_4	$\overset{+3}{Sb}_2O_3$
	(all oxides of Sn & Pb)	

d-block:

ZnO, $\overset{+3}{Cr}_2O_3$, $\overset{+4}{Mn}O_2$, $\overset{+5}{V}_2O_5$, $\overset{+4}{Ti}O_2$,

Respective Hydroxides of above shown oxides are also amphoteric.

[D]. Neutral oxides:

Oxides which do not react with acids or bases to form salt are neutral oxides.

Neutral oxides:

NO (Nitric oxide)

N_2O (Nitrous oxide)

CO (carbon monoxide)

H_2O

Diagonal relationship: few elements of second period resembles with diagonally placed third period elements. This is called diagonal relationship.

	Gr-1	Gr-2	Gr-13	Gr-14
2nd Period:	Li ↘	Be ↘	B ↘	C
3rd Period:	Na	Mg	Al	Si