$\frac{n}{p} > 1.5$

Classification Of Elements And Periodicity In Properties

Contributors — Destribution

(1) **John Newlands** – He was the 1st to correlate the chemical properties with their atomic masses. According to him if the elements are arranged in the order of their increasing atomic masses the eighth element starting from given one is similar in properties to the first one.
This arrangement of elements is called as Newland's law of Octave

Li^{7}	Be^{9}	B^{10}	C^{12}	N^{14}	O^{16}	F^{19}	Ne^{10}
Na^{23}	Mg^{24}	Al^{27}	Si^{28}	P^{31}	S^{32}	$Cl^{35.5}$	Ar^{40}
K^{39}	Ca^{40}	Cr^{52}	Ge^{72}	As^{75}	Se^{79}	Br^{80}	Kr^{84}

But this classification worked quite well for the lighter elements but it failed in case of heavier elements (upto Ca^{40}_{20}) and therefore discarded.

(2) **John Dobereiner's triads** – He arranged similar elements in the group of three elements called as triads, in which the atomic mass of the central element was merely the arithmetic mean of atomic masses of other 2 elements or all the three elements possessed nearly the same atomic masses.

Li^{7}	Na^{23}	K^{39}
Ca^{40}	Sr^{88}	Ba^{137}
$Cl^{35.5}$	Br^{80}	I^{127}
P^{31}	As^{75}	Sb^{122}

$\frac{7+39}{2} = 23$

$\frac{40+137}{2} = 88.5$

$\frac{35.5+127}{2} = 81$

$\frac{31+122}{2} = 76.5$

But this route wasn't followed by some triads, such as –

H^{1}	Li^{7}	Na^{23}
O^{16}	S^{32}	Se^{79}

$\frac{23+1}{2} = 12$?

$\frac{16+79}{2} = 47.5$?

Dimtri Mendeleev's periodic table

According to him the physical & chemical properties of the elements are the periodic functions of their atomic masses.

Property	eka-aluminium (predicted)	Gallium (found)	eka-silicon (predicted)	Germenium (found)
Atomic mass	68	70	72	72.6
Density (gm/cm^3)	5.9	5.94	5.5	5.36
Melting point (Kelvin)	Low	30.2	High	1231
Formula of Oxide	E_2O_3	Ga_2O_3	EO_2	GeO_2
Formula of Chloride	ECl_3	$GaCl_3$	ECl_4	$GeCl_4$

He Demonstrated a connection between atomic masses and elemental properties.

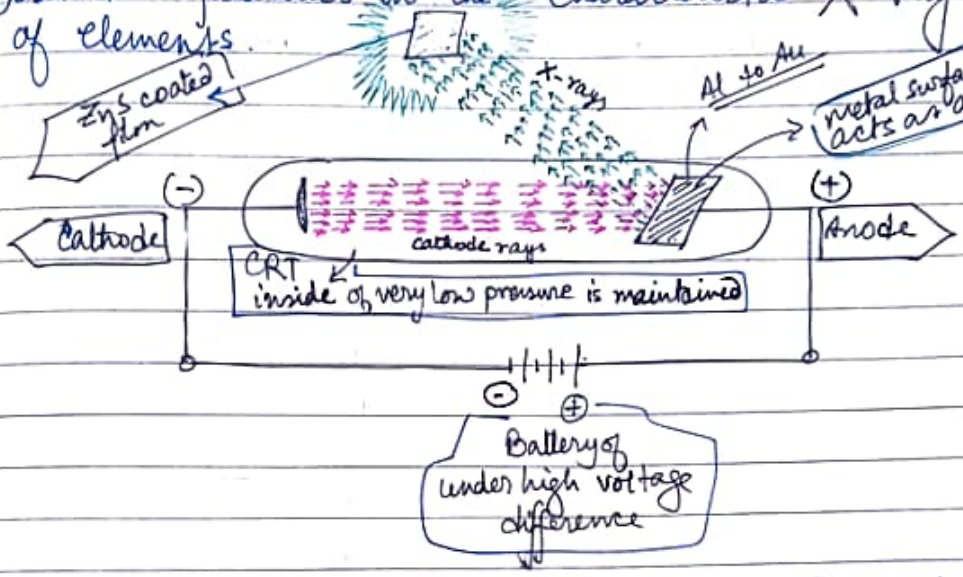
→ He arranged the elements by increasing atomic masses.

→ He predicted the existence & properties of undiscovered elements.

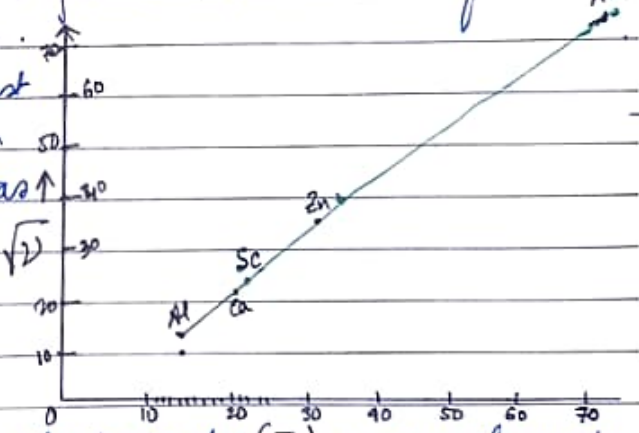
The organisation of one of Mendeleev's designs for periodic table

I	II	III	IV	V	VI	VII	VIII
H							
Li	Be	B	C	N	O	F	
Na	Mg	Al	Si	P	S	Cl	
K	Ca		Ti	V	Cr	Mn	Fe Ni Co Cu
Cu?	Zn			As	Se	Br	Ru Pd Rh Ag
Rb	Sr	yt	Zr	Nb	Mo		
Ag?	Cd	In	Sn	Sb	Te	I	
Cs	Ba		Ce				
		Ero	La	Ta	W		Os Pt Ir Au
Au?	Hg	Tl	Pb	Bi	U		
			Th				

Moseley's Law In 1913, Henry Moseley bombarded high speed electrons on different metal surfaces and obtained X-rays. He observed regularities in the characteristic X-ray spectra of elements.



X-rays are produced when solid anodes are bombarded with fast moving cathode rays. Moseley studied the X-ray spectra of elements lying between Al to Au in the periodic table and found the spectra followed the same sequence as in the periodic table. When he plotted $\sqrt{\nu}$ against atomic no. (Z) of elements, an excellent straight line was obtained but the relation between $\sqrt{\nu}$ against the atomic weights of the elements was not linear.

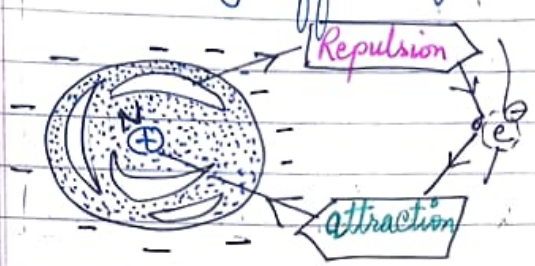


Then he concluded that atomic number (Z) of any element is more fundamental than its atomic weight.

The relation he deduced between the frequency (ν) of the characteristic X-ray line with Z (atomic no.) is given by $\sqrt{\nu} = a(Z-b)$ where Z and a, b are both constants for any particular type of line.

→ We know that the frequency of the emitted radiations as deduced from Bohr's theory is $\nu = \frac{2\pi^2 Z^2 e^4 m}{h^3}\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$

When this formula is applied to a multielectron system, the effective nuclear charge (Z_{eff}) shown by a given electron is not Ze but less than this due to shielding effect or screening effect of the inner electrons.

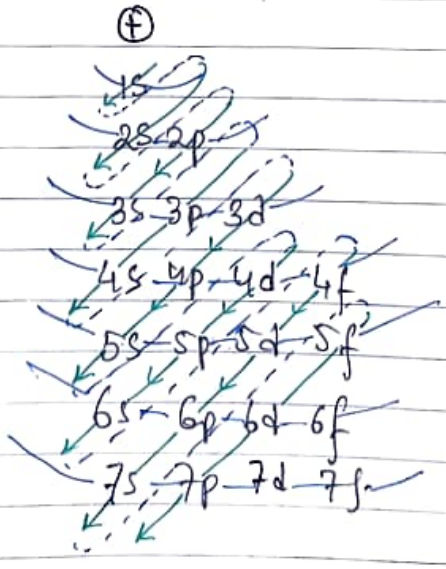


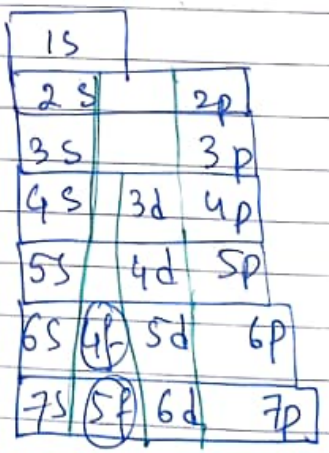
Screening/Shielding Effect

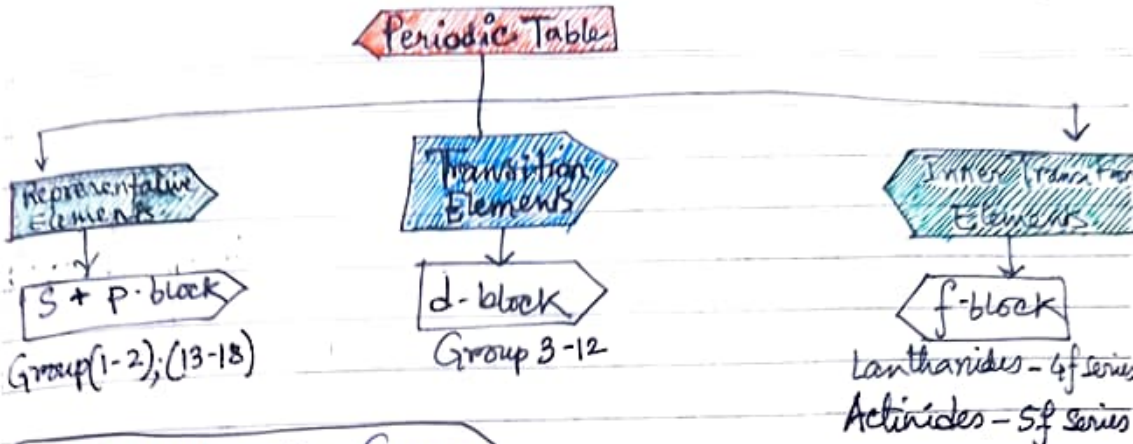
It's a combination of 2 opposite forces acted upon an $e^{\ominus}$ approaching towards an atom, the (+vely) charged nucleus exerts an electrostatic force of attraction on it which is opposed by the same charge - same charge repulsion applied by inner shell electrons in the atom, as a result of which the nucleus can't pull the approaching electron with its full charge rather a decreased amount of charge is applied on the electron. This is called as effective nuclear charge.

$$s > p > d > f$$

(Shielding/screening power increasing)







Vertical Columns – Groups

n th period – ns, np, (n-1)d, (n-2)f

Main

1S			1s
2S			2P
3S		Transition series	3P
4S		3d series	4P
5S	Inner Transition series	4d series	5P
6S	4f, Lanthanoids	5d series	6P
7S	5f, Actinoids	6d series	7P

① A new period always of its first election into the "s-orbital" of a new shell.

② In a period, shell number of p-orbital is same as that of s-orbital whereas the d-orbitals is one less and the f-orbitals are two less.

1S ⟶

2S ⟶ 2p

3S ⟶ 3p

4S ⟶ 3d → 4p

5S ⟶ 4d → 5p

6S → 4f ⟶ 5d → 6p

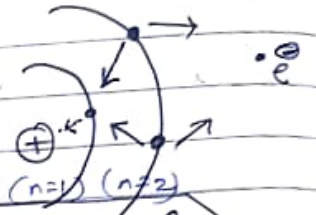
7S → 5f ⟶ 6d → 7p

The ↓ in the force of attraction exerted by the nucleus on the valence electrons due to the inner electrons is called Shielding effect.

$$Z_{effective} = Z_{normal} - \sigma$$ Where σ = Screening constant.

(Z_{eff} ↑ across a period)

	Li	Be	B	C	N	O	F	Ne
Z	3	4	5	6	7	8	9	10
Z_{eff}	1·28	1·91	2·42	3·14	3·83	4·45	5·10	5·76



Slater's Rules for Calculating σ or S

(1) All electrons in principle shells higher than the concerned electrons, contribute zero or 0.

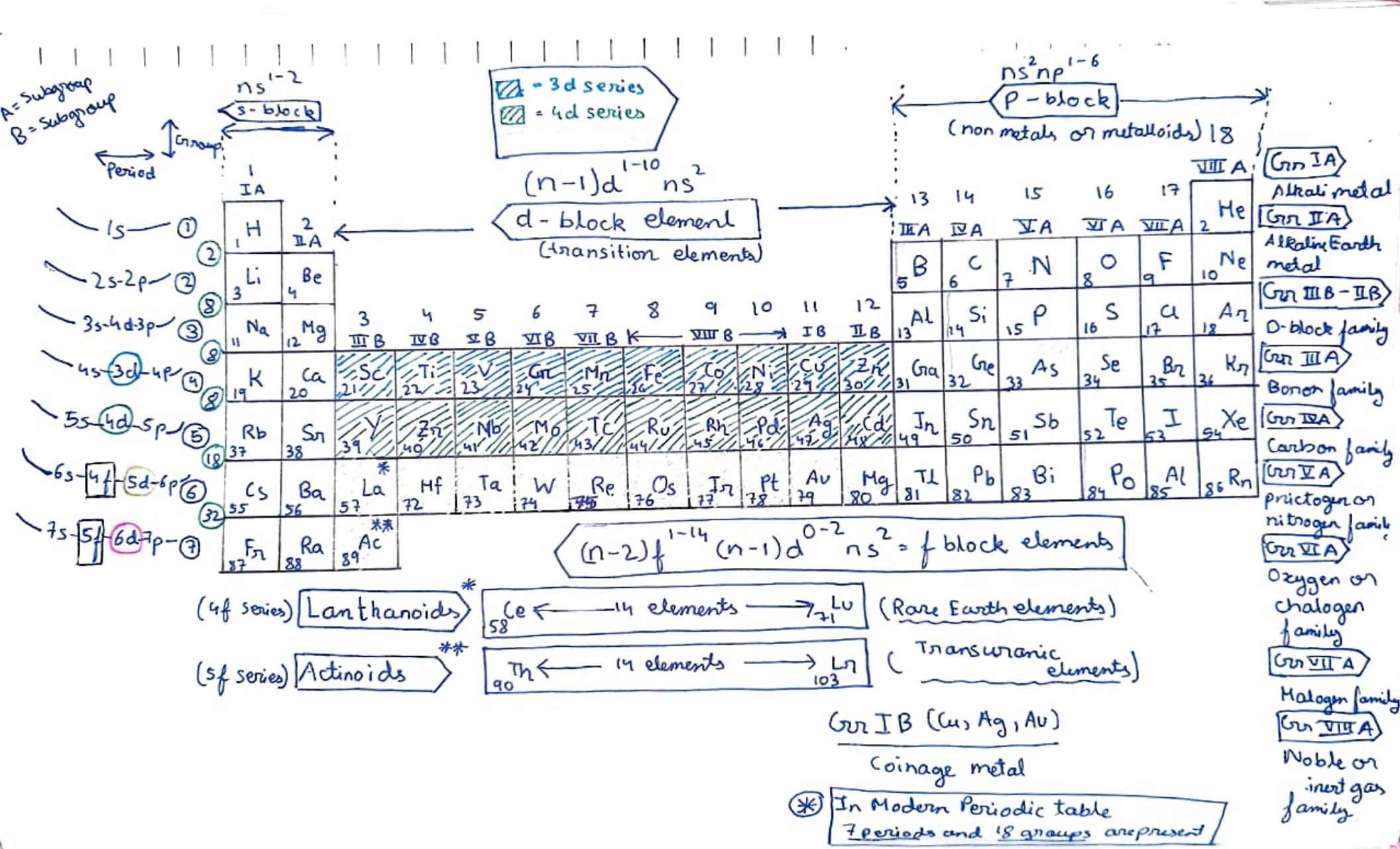
(2) Each $e^{\ominus}$ in the same principle shell contributes 0.35 to 0.

(3) Electrons in (n-1) shell each contributes 0.85 to 0 (for s or p-electrons)

(4) Electrons in deeper shell each contribute 1.00 to 0 (for s or p-electrons)

(5) Electrons in the (n-1) shell & lower shell each contribute 1.00 to 0 (for d or f-electrons)

Q. What is screening constant for outermost electron in Zn? (A) for a 4s $e^{\ominus}$ (B) For a 3d $e^{\ominus}$.

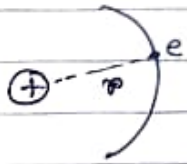

A = Subgroup
B = Subgroup
Period
Group
ns^{1-2}
s - block
= 3d series
= 4d series
$(n-1)d^{1-10}\ ns^2$
d - block element
(transition elements)
ns^2np^{1-6}
p - block
(non metals or metalloids) 18
1s
2s-2p
3s-4d-3p
4s-3d-4p
5s-4d-5p
6s-4f-5d-6p
7s-5f-6d-7p
$(n-2)f^{1-14}(n-1)d^{0-2}ns^2$ = f block elements
(4f series) Lanthanoids
58 Ce ← 14 elements → 71 Lu
(Rare Earth elements)
(5f series) Actinoids
90 Th ← 14 elements → 103 Lr
(Transuranic elements)
Grr IB (Cu, Ag, Au)
Coinage metal
In Modern Periodic table
7 periods and 18 groups are present
Grr IA Alkali metal
Grr IIA Alkaline Earth metal
Grr IIIB - IIB D-block family
Grr IIIA Boron family
Grr IVA Carbon family
Grr VA prictogen or nitrogen family
Grr VIA Oxygen or chalogen family
Grr VIIA Halogen family
Grr VIIIA Noble or inert gas family

b) $K \rightarrow [Ar]_{18} 4s^1$

eg $Cu \rightarrow [Ar]_{18} 4s^1 3d^{10}$

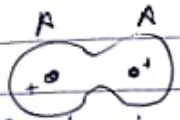
Periodic Trends in properties

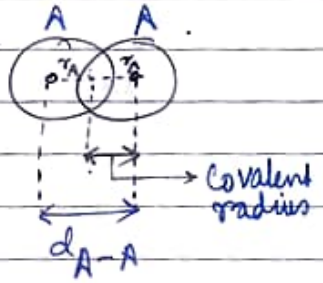
- **Atomic radius** :- It is the distance from the centre of the nucleus to the outermost electron or upto the point at which the probability of finding electron is the maximum.



Types of Atomic Radius

- **Covalent Radius** :- $r_{Covalent} = \frac{1}{2}$ [Bond length] → for homonuclear diatomic molecule like — Cl_2, Br_2, H_2, O_2, N_2 etc.

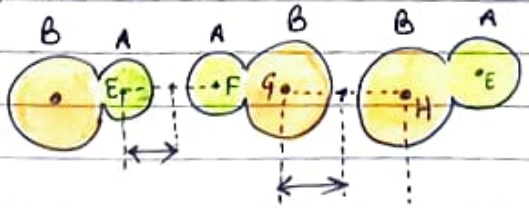




$d_{A-A} = r_A + r_A$

$r_A = \frac{d_{A-A}}{2}$ for homonuclear diatomic molecule

- **Vander Waal's Radius**: The sum of the VDW radii of 2 atoms is the shortest distance between 2 immediately adjacent atoms (either of same / different elements) in the structure of the solid compound in which they're not bonded to each other. eg- for noble gases like — He, Ne, Ar, Kr — etc.



$r_{VDW} = \frac{1}{2}$ [Distance between nuclei of 2 nonbonded atoms]

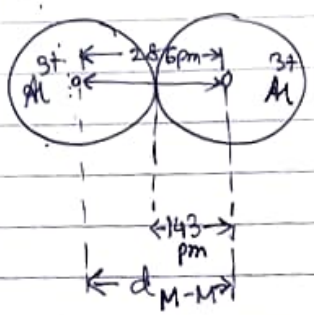
$\therefore \frac{1}{2}EF = r_{vanderwaals}$ of A in AB molecule

or $\frac{1}{2}GH = r_{vanderwaals}$ of B in AB molecule.

- **Metallic Radius** :- It's defined as half of the internuclear distance separating two adjacent metal atoms in a metallic lattice.

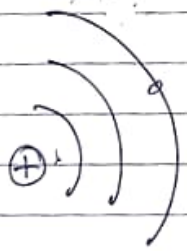
$$r_{metallic} = \frac{1}{2}\left[\text{Distance between 2 adjacent metal ions}\right]$$

Eg:



∴ 143 pm is the metallic radius of Al^{3+} ion.

- **Ionic Radius** : This is the distance between the nucleus and outermost orbit of an ion.



$$r_{ionic} \propto \frac{1}{Z_{effective}}$$

Z_{eff} = effective nuclear charge

Eg: The vanderwaal's radius of Chlorine atom is 1.80 Å
The covalent radius of Chlorine atom is 0.99 Å

- $r_{Vander\ Waals} > r_{metallic} > r_{covalent}$

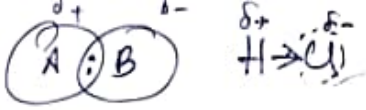
- $r_{Cation} < r_{parent\ atom} < r_{anion}$

- For isoelectronic (having same no. of electrons) species such as -

$$N_3^{3-} > O^{2-} > Na^{+} > Mg^{2+}$$

← Atomic size increases

$$\text{Size of Atom} \propto \frac{1}{\text{atomic number (Z)}} \text{ or } \frac{1}{(Z/e)\text{ value}}$$



However if atoms forming the covalent bond are different, i.e., one atom is more electronegative than the other, then the atomic radius is determined by the relation :-

for heteronuclear diatomic molecule AB

$$d_{A-B} = r_A + r_B - 0.09(x_A - x_B)$$

Where x_A & x_B are electronegativities of the Atoms A & B respectively. This is known as Schomaker-Stevenson equation in 1941.

11/08/22

S-Block Elements

classification based on electronic Configuration
- If the last electron enters into "S"-orbital

① General electronic configuration of S-block elements = ns^{1-2}

Alkali metals all have valency = 1
Alkaline Earth metals " " valency = 2

③ Reactive metals with low ionisation energies

S-Block

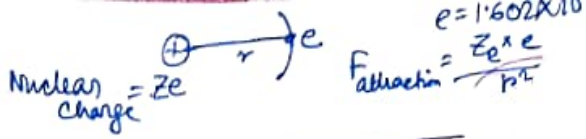
ns^1	ns^2
H	
3 Li	4 Be
11 Na	12 Mg
19 K	20 Ca
37 Rb	38 Sr
55 Cs	56 Ba
87 Fr	88 Ra

P-Block Elements

- Last electron enters into the p-orbital.
- General valence shell electronic configuration is :-

13	14	15	16	17	18
ns^2np^1	ns^2np^2	ns^2np^3	ns^2np^4	ns^2np^5	ns^2np^6

- Valency of noble gases is zero.
- Nonmetallic nature increases across the period.
- Metallic nature increases down the group.



Some Exceptions in electronic Configurations

Traditionally this phenomenon has often been explained as a consequence of the, "special stability of half-filled subshells".

A more complete explanation considers both the effects of increasing nuclear charge on the energies of the 4s & 3d levels & the interactions between electrons sharing the same orbital. This approach requires totaling all contributions to the energy of the configurations of electrons including the Coulombic and exchange energies; results of the complete calculations are determined by experimental data.

Some Exceptions in Electronic Configurations	Expected	Actual
Cr (Z=24)	$[Ar]_{18}\,4s^2 3d^4$	$[Ar]\,4s^1 3d^5$
Cu (Z=29)	$[Ar]_{18}\,4s^2 3d^9$	$[Ar]\,4s^1 3d^{10}$
Nb (Z=41)	$[Kr]_{36}\,5s^2 4d^3$	$[Kr]_{36}\,5s^1 4d^4$
Mo (Z=42)	$[Kr]_{36}\,5s^2 4d^4$	$[Kr]_{36}\,5s^1 4d^5$
Ru (Z=44)	$[Kr]_{36}\,5s^2 4d^6$	$[Kr]_{36}\,5s^1 4d^7$
Rh (Z=45)	$[Kr]_{36}\,5s^2 4d^7$	$[Kr]_{36}\,5s^1 4d^8$
Pd (Z=46)	$[Kr]_{36}\,5s^2 4d^8$	$[Kr]_{36}\,5s^0 4d^{10}$
Ag (Z=47)	$[Kr]_{36}\,5s^2 4d^9$	$[Kr]_{36}\,5s^1 4d^{10}$
La (Z=57)	$[Xe]_{54}\,6s^2 4f^1$	$[Xe]_{54}\,6s^2 5d^1$
Ce (Z=58)	$[Xe]_{54}\,6s^2 4f^2$	$[Xe]_{54}\,6s^2 4f^1 5d^1$

Variation of atomic size

$r_n \propto \frac{1}{Z_{eff}}$

$r_n = \frac{n^2}{Z} a_0$

a_0 = Bohr's 1st orbit = 0.53 Å

Atomic size decreases across a period, Z_{eff} increases across a period.

Element	$_3Li$	$_4Be$	$_5B$	$_6C$	$_7N$	$_8O$	$_9F$
Size in (pm)	152	111	88	77	74	66	64
Element	$_{11}Na$	$_{12}Mg$	$_{13}Al$	Si	P	S	Cl
Size (pm)	186	160	143	117	110	104	99

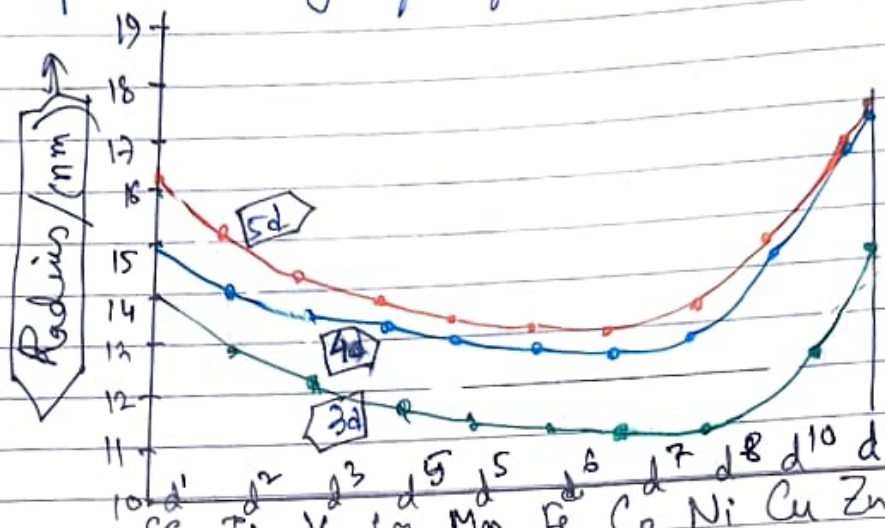
✱ for d-block elements, the atomic radius initially decreases, then remains constant and finally increases again.

- Atomic size increases generally down the group due to addition of new shell.

Atom (Group I)	Atomic Radius (pm)	Atom (Group 17)	Atomic Radius (pm)
Li	152	F	72
Na	186	Cl	99
K	231	Br	114
Rb	244	I	133
Cs	262	At	140

- Inert gases have the largest atomic size in their period because we define V.D.W radii for them.
- For d-block elements the trend in atomic radii is r_{3d} series < r_{4d} series ≃ r_{5d} series.
- The radius increases as the shell increases. The radius of elements of 4d series is about the same as

elements of 5d series. This is due to the lanthan contraction that causes increase in Z_{eff} value due to poor shielding of 14f $e^{\ominus}$s.



The re result contra is that 4d and series simila radii (Zn - 160 Hf - 15 and ha very s physic & che properti much than the expecte

series
3d ——
4d —— Y Zr Nb Mo Tc Ru Rh Pd Ag Cd
5d —— La Hf Ta W Re Os Ir Pt Au Hg

Triad group.

Extent of shielding $s > p > d > f$

The lanthanoid contraction is greater than expected decrease in ionic radii of the elements in the lanthanide series from atomic number 57, lanthanum to 71, lutetium which results in smaller than otherwise expected ionic radii for the subsequent elements starting with 72, hafnium.

Lanthanoid Contraction

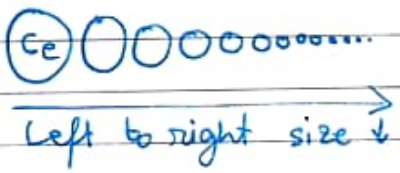
4d series	Zr radius	0.162 nm
5d series	Hf "	0.160 nm

It is evident from the general electronic configuration of f-block elements

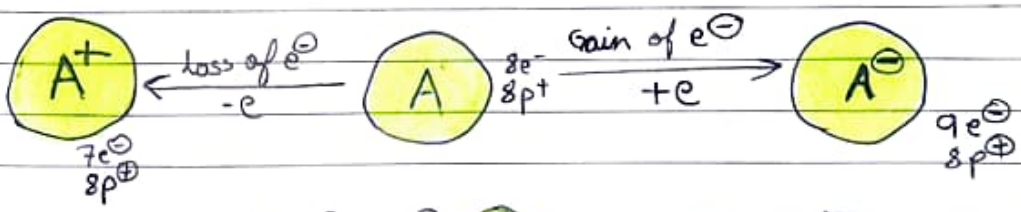
i.e $(n-2)f^{1-14}\ (n-1)d^{0-2}\ ns^{2}$ —

due to poor screening effect or shielding effect of f-orbital along with d-orbital. The increased nuclear charge pulls the outermost s-orbital two electrons much effectively that the radius of the elements contract towards the centre. As a result of which - a regular ↓ in size of Lanthanoid elements will take place which is known as Lanthanoid contraction.

Actinoid contraction is more effective than Lanthanoid contraction



Ionic Radius Formation of +ve & -ve ions from an atom is shown below



$A^{3-} > A^{2-} > A^{\ominus} > A > A^{+} > A^{2+} > A^{3+}$

The order of radius is: $A^{\ominus} > A > A^{\oplus}$

Na	Na^+
186	102
Mg	Mg^{2+}
160	72
Al	Al^{3+}
143	54
N	N^{3-}
75	146
O	O^{2-}
73	140
F	$F^{\ominus}$
72	133

Exceptions

- In every and each period, noble gases have largest in size (largest radius) due to electronic repulsion is dominating i.e. completely electron filled.
- For transition elements (d-block element) we are not observing regular trends
- B < Ga < Al < In = Tl

88 < (135 < 143) < 167 ≈ 170 pm

(?) (?)

Gr 13

B
Al
Ga
In
Tl

↓ down the group.

Li^+	Be^{2+}
76	45
Na^+	Mg^{2+}
102	72
K^+	Ca^{2+}
138	100
Rb^+	Sr^{2+}
152	118
Cs^+	Ba^{2+}
167	135

Element Gr-13	Atomic no. (Z)	Electronic Configuration
Boron	5	$[He]_2 2s^2 2p^1$
Aluminium	13	$[Ne]_{10} 3s^2 3p^1 (3d^0)$
Gallium	31	$[Ar]_{18} 4s^2 (3d^{10}) 4p^1$
Indium	49	$[Kr]_{36} 4d^{10} 5s^2 5p^1$
Thallium	81	$[Xe]_{54} (4f^{14}) 5d^{10} 6s^2 6p^1$

N^{3-}	O^{2-}	F^-
146	140	133
P^{3-}	S^{2-}	Cl^-
212	184	181

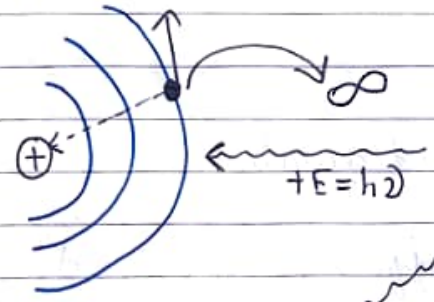
(?) This is due to poor Shielding effect of d-orbital in Ga.

(?) This is due to poorest Shielding effect of f-orbital in Tl.

Periodic Properties

Ionization Energy / Enthalpy.

It's the minimum amount of energy required to remove the most losely bound electron from the valence or outermost shell of an isolated, neutral, gaseous atom in its ground energy state to the infinity or zero energy state.



$$M_{(g)} - e^{\ominus} \xrightleftharpoons{(I.E)_1} M^{\oplus}_{(g)}$$

Isolated neutral atom

$$M^{\oplus}_{(g)} - e^{\ominus} \xrightleftharpoons{(I.E)_2} M^{2+}_{(g)}$$

$$M^{2+}_{(g)} - e^{\ominus} \xrightleftharpoons{(I.E)_3} M^{3+}_{(g)}$$

$$(I.E)_3 > (I.E)_2 > (I.E)_1$$

Energy increasing.

Factors affecting Ionisation Energy

- Effective Atomic number, $Z_{eff} \propto I.E$
- Atomic size $\propto \frac{1}{I.E}$
- Charge on ion → $\uparrow \oplus \propto (IE)\uparrow$
  $\uparrow \ominus \propto \frac{1}{(I.E)} \downarrow$
- Electronic configuration.
- Penetration effect, $s > p > d > f$
  ← penetrating tendency
- Shielding effect $\propto I.E$

Variations of Shielding effect

Along the 2nd period

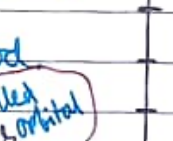
$_3Li \rightarrow 1s^2 2s^1$ ① half filled 2s orbital

$_4Be \rightarrow 1s^2 2s^2$ ② fully filled 2s orbital

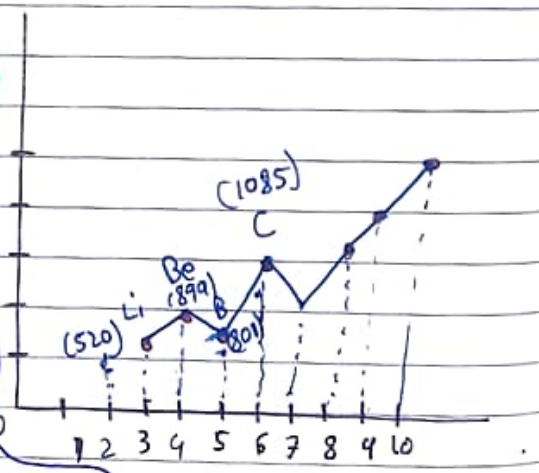
$_5B \rightarrow 1s^2 2s^2 2p^1$ incompletely filled p-orbital

$_6C \rightarrow 1s^2 2s^2 2p^2$

$_7N \rightarrow 1s^2 2s^2 2p^3$



Half filled p orbital is more stable than incompletely filled p-orbital



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- Electronic configuration.
- Penetration effect, $s > p > d > f$

- **Shielding effect**

Variations of Sh effect

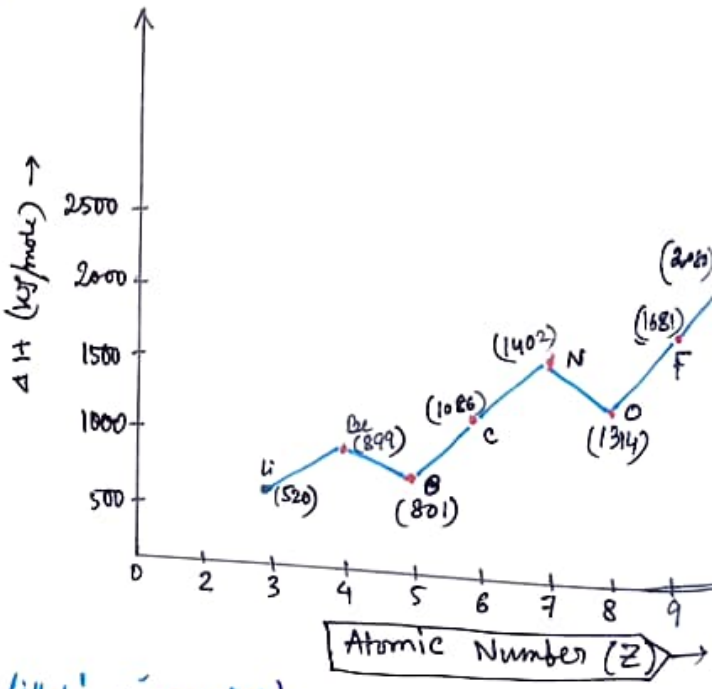
Along the 2nd pe

$_3Li \rightarrow 1s^2 2s^1$ (1)

$_4Be \rightarrow 1s^2 2s^2$ (2)

$_5B \rightarrow (1s^2 2s^2 2p^1)$

filled 's'-orbital



$_6C \rightarrow 1s^2 2s^2 2p^2$

$_7N \rightarrow 1s^2 2s^2 (2p^3)$

$+\frac{1}{2}$ $+\frac{1}{2}$ $+\frac{1}{2}$

$2p_x$ $2p_y$ $2p_z$

Half filled orbital is more stable than incompletely filled p-orbital

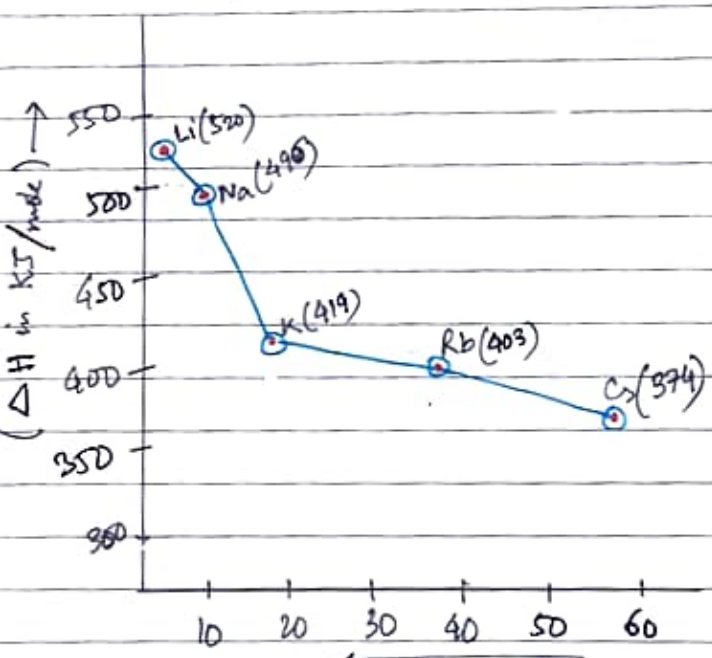
$_8O \rightarrow 1s^2\ 2s^2\ 2p^4$ | $2p_x$ | $2p_y$ | $2p_z$ → incompletely filled p-orbital

$_9F \rightarrow 1s^2\ 2s^2\ 2p^5$ | $2p_x$ | $2p_y$ | $2p_z$ → " "

$_{10}Ne \rightarrow 1s^2\ 2s^2\ 2p^6$ | $2p_x$ | $2p_y$ | $2p_z$ → fully filled 2p orbital

Variation of I.E of elements down the group.

I.E ↓ down the group normally.



Electron Gain Enthalpy or Electron Affinity

ns^2np^3 VA | ns^2np^4 VIA | ns^2np^5 VIIA | ns^2np^6 VIIIA | ns^1 IA | ns^2 IIA | ns^2np^1 IIIA

Nitrogen family | Oxygen family | Halogen family | Gaining of electron | NOBLE GAS Family | Alkali metal | Alkaline Earth metal | Boron family | losing of

+e | +2e | +3e | -e | -2e

E. Affinity (ΔH_{eg} or E.A) → electron gain enthalpy

It's the energy released when one mole of $e^{\ominus}$ is added to one mole of an element in its isolated gaseous state or ground energy or lowest energy state.

- $Li_{(g)} + e^{\ominus} \longrightarrow Li^{\ominus}_{(g)}$; $\Delta H_{eg} = -57$ Kj/mole

  $H_{(g)} + e^{\ominus} \longrightarrow H^{\ominus}_{(g)}$; $\Delta H_{eg} = -72$ kj/mole

- $(E.A)_2$ is always +ve → because of electrostatic repulsion between anion and electron (having same charge)

$$A_{(g)} + e^{\ominus}_{(g)} \xrightarrow[\text{Exothermic}]{(E.A)_1} A^{\ominus}_{(g)} \xrightarrow[\text{endothermic}]{+e^{\ominus},\ +(E.A)_2} A^{2-}_{(g)}$$

Electron Affinity depends upon the following factors :-

- It decreases when radius increases. However, this is having the following exceptions.

  ΔH_{eg} of Cl > ΔH_{eg} of F (?) Explain why

  ΔH_{eg} of S > ΔH_{eg} of O (?)

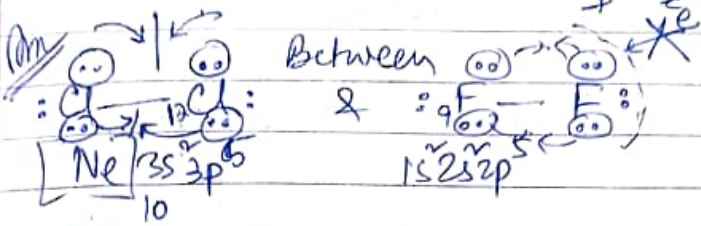
  ΔH_{eg} of P > ΔH_{eg} of N (?)

- It increases ↑ across a period when Z_{eff} ↑.
- E.A of half filled & fully filled electronic configuration is less (mostly endothermic)
- E.A order for addition of $e^{\ominus}$ in different orbitals having the same principal quantum number is s > p > d > f.

$$E.A \propto \frac{1}{\text{atomic size}}$$

Why does Electron affinity of Cl is greater than that of Fluorine ?

Why does the value of E.A of Cl is more negative than that of F ?



Size of F is very compact compare than Cl. lone pair - lone pair repulsion in Fluorine becomes more effective than that of Chlorine, making the bond dissociation energy of F_2 molecule weaker than that of Cl – Cl.

But in Chlorine due to availability of 3d vacant orbitals it can accomodate an extra electron in it, since Electron gain enthalpy in Chlorine > E.g of Fluorine.

- Oxygen & fluorine have no extra space or an empty d-orbital due to smaller atomic size they can't accomodate, / inter electronic repulsion → becomes greater, but in Sulphur & Chlorine both have vacant 3d orbitals since they can accomodate extra electrons.

▶ How can we understand the valency of a metal from its (I E) value

Element	$(I.E)_1$	$(I.E)_2$	$(I.E)_3$	$(I.E)_4$	$(I.E)_5$	$(I.E)_6$	$(I.E)_7$	
Na	496	4560						valency = 1
Mg	738	1450	7730					" = 2
Al	578	1820	2750	11600				" = 3
Si	786	1580	3230	4360	16100			" = 4
P	1012	1900	2910	4960	6270	22200		= 5
S	1000	2250	3360	4560	7010	8500	27100	4, 2, 6
Cl	1251	2300	3820	5160	6540	9460	11000	1-7
Ar	1521	2670	3930	5770	7240	8780	12000	0

Q. IE_1, $(I.E)_2$, $(I.E)_3$ for an element are 7 eV, 15 eV and 42.5 eV respectively. The most stable oxidation state of the element will be 3.

Gr I	$\Delta_{eg}H$	Gr-16	$\Delta_{eg}H$	Grr-17	$\Delta_{eg}H$	Group 0	ΔH_{eg}
H	-73					He	+48
Li	-60	Oxygen	-141	F	-328	Ne	+116
Na	-53	Sulphur	-200	Cl	-349	Ar	+96
K	-48	Selenium	-195	Br	-325	Kr	+96
Rb	-47	Tellurium	-190	I	-295	Xe	+77
Cs	-46	Polonium	-174	At	-270	Rn	+68

EA ↓ down the group but accross the period EA. ↑ as atomic size ↓.

Electronegativity It's the tendency of the atom to attract electrons (bond pair) towards itself when combined in a compound.

- Electronegativity for an individual atom in a nonbonded state doesn't have any meaning.
- Electronegativity depends on bonding conditions.
- Electronegativity of any given element isn't constant; it varies depending on the element to which it's bound.

Periodic Trends of E.A (E.A = electron affinity)

Process	$\approx \Delta_{E.A}H$ (kJ/mole)
$H_{(g)} + e^{\ominus} \longrightarrow H^{\ominus}$	-73
$Li_{(g)} + e^{\ominus} \longrightarrow Li^{\ominus}_{(g)}$	-60
$Na_{(g)} + e^{\ominus} \longrightarrow Na^{\ominus}_{(g)}$	-53
$K_{(g)} + e^{\ominus} \longrightarrow K^{\ominus}_{(g)}$	-48
$N_{(g)} + e^{\ominus} \longrightarrow N^{\ominus}_{(g)}$	$\approx$ "0" ?
$P_{(g)} + e^{\ominus} \longrightarrow P^{\ominus}_{(g)}$	-72 ?

~~Periodic~~ Periodic Trends of E.A

	$\approx \Delta H_{EA}$ (kJ/mole)
$O_{(g)} + e^{\ominus} \longrightarrow O^{\ominus}_{(g)}$	−141
$O^{\ominus}_{(g)} + e^{\ominus} \longrightarrow O^{2-}_{(g)}$	+798
$S_{(g)} + e^{\ominus} \rightarrow S^{\ominus}_{(g)}$	−201
$S^{\ominus}_{(g)} + e^{\ominus} \rightarrow S^{2-}_{(g)}$	+640
$F_{(g)} + e^{\ominus} \longrightarrow F^{\ominus}_{(g)}$	−328
$Cl_{(g)} + e^{\ominus} \longrightarrow Cl^{\ominus}_{(g)}$	−349
$Br_{(g)} + e^{\ominus} \longrightarrow Br^{\ominus}_{(g)}$	−325
$I_{(g)} + e^{\ominus} \longrightarrow I^{\ominus}_{(g)}$	−295.

In a group There's a decrease (↓) of electroaffinity in general, but for "p-block" elements only the (E.A)$_1$ of second period element is less than (E.A)$_1$ of the third period elements.

2nd period		3rd Period	E.A value of noble gas family
			He +20
B −23	<	Al −44	Ne −30
C −120	<	Si −123	Ar +35
"0" N	<	P −72	Kr +40
O −141	<	S −201	Xe +40
F −322	<	Cl −348	Rn +40

Along a Period The $e^{\ominus}$ affinity (↑), increases from left to right as Z_{eff}, effective nuclear charge (↑) increases (except inert gases). Some exceptions to this trend are :-

(i) (E.A) of alkali metals > E.A of Alkaline Earth Metals.

(ii) (E.A) of Carbon > E.A of Nitrogen

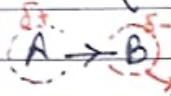
B.E = Bond dissociation Energy

Electronegativity The electronegativity of an element is the power of an atom of the element to attract the bonded pair of electrons within a molecule.

Pauling Scale for Electronegativity Linus Pauling developed a method for calculating relative electronegativities of most elements. According to Pauling scale :

$$\Delta x = x_A - x_B = 0.208\sqrt{B.E_{A-B} - \sqrt{B.E_{A-A} \times B.E_{B-B}}}$$

Where
- E_{AB} = Bond enthalpy/Bond energy of A-B bond.
- E_{A-A} = Bond enthalpy/Bond energy of A-A bond
- E_{BB} = Bond enthalpy/Bond energy of B-B bond.



if B is more Electronegative than A

then Polarity develops

Element	Pauling's Electronegativity
F	3.98
O	3.44
Cl	3.16
N	3.04
S	2.58
C	2.55
H	2.20
P	2.19
Na	0.93
Fr	0.70

NCERT

Atom (Period II)	Li	Be	B	C	N	O	F
Electronegativity	1.0	1.5	2.0	2.5	(3.0)	3.5	4.0

Atom (Period III)	Na	Mg	Al	Si	P	S	Cl
Electronegativity	0.9	1.2	1.5	1.8	2.1	2.5	(3.0)

Atom (Group I)	E.N. values
Li	1.0
Na	0.9
K	0.8
Rb	0.8
Cs	0.7

↑ E.N

across the period

down the group

E.N ↓

∴ $E.N \propto \frac{1}{\text{atomic size}}$

Mulliken's Scale

Electronegativity, χ (Chi) can be regarded as the average of the ionisation energy (I.E) and the electron affinity (E.A) of an atom (both expressed in electronVolts)

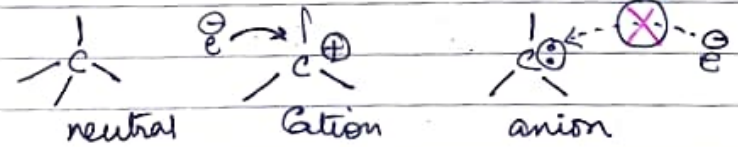
$$\chi_{(M)} = \frac{I.E + E.A}{2}$$

Mulliken

Mulliken's values were about 2.8 times larger than the Pauling's values.

Electron Gain Enthalpy	Electronegativity
▶ It's the tendency of an atom to attract outside electrons.	▶ It's the tendency of an atom to attract shared pair of electron.
▶ It's the property of an isolated atom.	▶ It's the property of the bonded atom.
▶ It's the absolute $e^{\ominus}$ attracting tendency of the atom.	▶ It's the relative $e^{\ominus}$ attracting tendency of an atom.
▶ It has units like - KJ/mole and eV/atom.	▶ It has no units

- Charge on cationic species increases ↑, electronegativity ↑
- Charge on anionic species increases ↑, electronegativity ↓



Application of Electronegativity

1. Nomenclature → Such as, in NaCl, first name is of cation & then the name of anion
2. Metallic and nonmetallic Nature
3. Partial ionic character in covalent bonds.

if difference of electronegativities of the 2 elements is 1.7 or more then ionic bond is formed between them, whereas if it's less than 1.7 then covalent bond is formed. HF is exception in which bond is covalent although the difference of electronegativity

$$\% \text{ of ionic character} = 16|\Delta X| + 3.5(\Delta X)^2$$

Bond Strength & Stability As the difference in the electronegativities between 2 bonded atoms in a molecule increases the bond strength and stability increases.
For example: $H-F > H-Cl > H-Br > H-I$.

The acidic/basic properties of metal oxides can be predicted. As Δ E.N increases (↑), the acidic/basic properties of metal oxides can be predicted. As Δ E.N increases (↑), the acidic strength of oxides of different elements decreases (↓).

(i) $Li_2O < BeO < B_2O_3 < CO_2 < N_2O_5$
Basic Amphoteric Acidic

(ii) $Na_2O < MgO < Al_2O_3 < SiO_2 < P_2O_5 < SO_3 < Cl_2O_7$
Basic Amphoteric acidic

(iii) $\overset{+1}{N_2O} < \overset{+2}{NO} < \overset{+3}{N_2O_3} < \overset{+4}{NO_2} < \overset{+5}{N_2O_5}$
Neutral Acidic

(iv) $\overset{+2}{CO} < \overset{+4}{CO_2}$ (v) $\overset{+4}{SO_2} < \overset{+6}{SO_3}$

(vi) $\overset{+4}{CO_2} > \overset{+4}{SiO_2}$
Smaller size → Larger size