Periodic Table

1. Development

① Lavoisier classification
- metals → those elements which have the tendency to lose e^-s.
- non-metals → tendency of gaining e^-s.

② Prout's Hypothesis (Unitary Method)
→ he simply assumed that all elements are made up of hydrogen.

$$\text{Atomic wt. of element} = n \times (\text{At. wt. of 1 H atom})$$

[At. wt of H = 1]

③ Dobereiner Triad Rule
→ groups of 3 elements having similar chemical properties.
- At. wt. of the middle element is nearly equal to the avg. at. wt. of first and third element.

Eg - (K, Rb, Cs) (H, F, Cl) (P, As, Sb) (S, Se, Te).

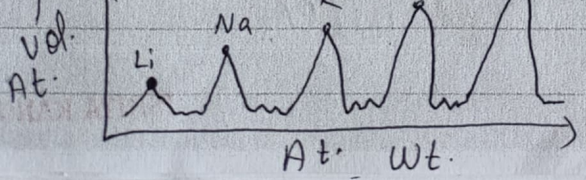
④ Newland Octave Rule
→ he arranged elements in increasing order of their atomic masses and observed that properties of every 8th element was similar to the 1st element (like in the case of musical vowels notation).

Li	Be	B	C	N	O		F
Na	Mg	Al	Si	P	S		Cl
K	Ca						

H

Limitation ⇒ this was applicable only till Ca due to presence of d-block.

⑤ Lothar Meyer's Curve
→ he plotted a curve b/w at. wt. and at. volume of elements.
(a) Most electro +ve (alkali metals) were at the peaks
(b) Less electro +ve (alkaline earths) occupied the descending positions.
(c) Metalloids and transition elements occupied the bottom part of the curve.
(d) Most electro -ve (Chalogens) occupied the ascending positions of curve.



PT

⑥ Mendeleev's Periodic Table

→ The physical and chemical properties of elements are the periodic functions of their atomic weight.

- 63 elements
- no noble gases
- 7 periods, 8 groups
- elements belonging to the same group exhibit similar properties.
- Gaps were left in the PT
  - Eka Al = Ga
  - Eka B = Sc
  - Eka Si = Ge
  - Eka Mn = Tc.
- No position of H
- position of isotopes not defined
- like elements were in diff. groups - Pt, Au
- unlike elements were in the same group - Na, K, Cu, Hg, Au
- not clear that lanthanides, actinides are related to IIIA or III B group.
- anomalous pairs of elements $\begin{pmatrix} Ar & K \\ 39.9 & 39.1 \end{pmatrix}$ $\begin{pmatrix} Co & Ni \\ 58.9 & 58.6 \end{pmatrix}$

⑦ Modern Periodic Table (Moseley)

→ The physical and chemical properties of elements are the periodic functions of their at. no.

→ $\sqrt{\nu} \propto Z$; ν = frequency of x-rays, Z = at. no.

→ 9 groups, 7 periods, Inert gases were introduced by Ramsay.

⑧ Long Form of PT / Bohr-Bury-Rang.-Werner PT

→ based on Bohr-Bury electronic config. concept

→ introduced by Rang and Werner.

→ 7 periods, 18 groups

101 = Unnilunium
102 = Unnilbium
111 = Ununium
110 = Ununnilium
109 = Unnilennium

2. Some Important Points

→ 2nd period elements (Li, Be, B) show diagonal relationship with 3rd period elements (Mg, Al, Si).

→ 3rd period elements except inert gases are called typical elements.

→ Elements after Z = 92 (U) are called transuranic elements. They are radioactive and artificial.
First Man-made element = Tc
First man-made lanthanoid = Pm.

→ In the PT - No. of gases = 11 No. of liquids = 6
No. of solids = 95

III B group is the longest having 32 elements including 28 f-block elements.

→ In the PT, elements with similar properties occur at intervals of 2, 8, 8, 18, 18, 32. These numbers are called magic numbers.

3. Screening Effect (σ) and Effective Nuclear Charge (Z_{eff})

$Z_{eff} = Z - \sigma$ (Z = nuclear charge)

→ σ for outermost orbit (ns, np) = 0.35
σ for penultimate orbit (n-1) = 0.85
σ for (n-2) orbit and below = 1.
Except for He ($1s^2$)
σ of one 1s e^- = 0.30 ∴ $Z_{eff} = 2 - 0.3 = 1.7$.
From top to bottom in a group, Z_{eff} remains constant.

→ For same shell, shielding / screening effect
s > p > d > f.

→ Z_{eff} for different ions of an element - $N^+ > N > N^-$.

→ For isoelectronic species, Z_{eff} - $H^- < Li^+ < Be^{+2} < B^{+3}$

$Z_{eff} \propto \frac{\text{positive charge}}{\text{negative charge}}$

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4. Atomic Radius

(a) covalent radius (b) Ionic radius
(c) Metallic radius (d) vander waal's radius.

① **Covalent radius**

$d_{A-A} = 2r_A$.

② **Ionic radius** – $r^- > r > r^+$

→ For isoelectronic species

$N^{3-} > O^{2-} > F^- > Na^+ > Mg^{+2} > Al^{+3}$.

③ **Metallic radius** $\propto \frac{1}{\text{Metallic bond strength}}$

④ **Van der waal's Radius** – Inert gases only have vander waal radius.

→ In molecules of solid non-metals, both covalent and van der waal's radius exists.

Vander waal's radius $\approx 2\times$ covalent radius.

Vander waal's radius > Metallic Radius > Covalent Radius.

→ Along the period, radius decreases.
Down the group, radius increases.

→ **Transition elements** –

$Sc > Ti > V > Cr < Mn > Fe \approx Co \approx Ni \geq Cu < Zn$.

→ In f-block, along the period, radius ↓es continuously due to lanthanoid and actinoid contraction.

→ **Group 4-11** – $3d < 4d \approx 5d$ chemical twins } d-block

Group 3 – $3d < 4d < 5d \approx 6d$.

5. Ionisation Potential/Energy

minimum energy required to remove an e^- from the outer most shell of an isolated gaseous atom.

$E_1 < E_2 < E_3$ (always).

$\Delta H = +ve$ (endo)

→ half-filled and fully-filled orbitals are more stable.

→ **IP $\propto$ Z_{eff}.**

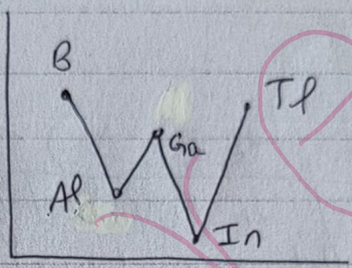
→ $IP \propto \frac{1}{\text{Atomic size}}$

→ Along a period, IP increases.
Down the group, IP decreases.

→ In d- block
Group 3 - $3d > 4d > 5d$
group 4, 5, 6, 10 - $3d < 4d < 5d$
Group 7, 8, 9, 11, 12 - $4d < 3d < 5d$.

→ IP for B family

∴ B > Tl > Ga > Al > In



→ Metallic character $\propto \frac{1}{IE} \propto$ Reactivity of metal.

→ If the difference b/w 2 successive IE $\geq$ 16 eV ⇒ lower OS is more stable.

If the difference b/w 2 successive IE $\leq$ 11 eV ⇒ higher OS is more stable.

→ To determine the no. of valence e⁻s ⇒
No. of valence e⁻s = number of lower values of IP before 1st highest jump.

6. Electron Affinity / Electron Gain Enthalpy
amount of energy released when an atom in the gaseous state accepts an e⁻ to form an anion.

→ Generally EA_1 = -ve (exo)
Always EA_2 = +ve (endo).

→ $X_{(g)} + e^- \underset{IP}{\overset{EA}{\rightleftharpoons}} X^-_{(g)}$

→ $EA \propto Z_{eff} \propto \frac{\text{Positive charge}}{\text{Negative charge}} \propto \frac{1}{\text{Atomic size}}$

→ Along the period, EA ↑es
→ Down the group, EA ↓es.

EA of elements having half-filled and fully-filled config is less or zero. ∴ for those elements ΔH = +ve (endo). Eg – Be, N, Ne, Mg.
∴ for Li, B, C, Be, N, O, F, Ne
EA Ne < B < N < Li < C < O < F
endo

7. Electronegativity (EN) – explained by Pauling for the 1st time.
→ the tendency of a covalently bonded atom to attract shared pair of e⁻s towards itself is EN.

EN
C = S = I = 2.5
N = Cl = 3
P = H = 2.1
F = 4
O = 3.5
K = Rb = 0.8
Cs = Fr = 0.7
Na = 0.9
Li = 1.0

Small atoms normally have more EN than large atoms.

→ EN – Along the period ↑es
Down the group ↓es.
→ EN ∝ Z_{eff} ∝ $\frac{\text{positive charge}}{\text{negative charge}}$ ∝ $\frac{1}{\text{Atomic size}}$

EA
Cl > F

EN
F > Cl.

EN Ga > Al (exception)

① Nature of hydroxide -

→ XOH → XO^- H^+ (acidic)
→ XOH → X^+ OH^- (basic)

As per Gallis, % EN of X is > 1.7 ⇒ acidic
% EN of X is < 1.7 ⇒ basic

② Nature of oxide -

→ Acidic strength of oxide and oxyacid ∝ EN.
→ Acidic nature ∝ OS

Except - H_3PO_2 (+1) > H_3PO_3 (+3) > H_3PO_4 (+5)
acidic strength

③ Nature of bonds -

Acc. to Hanny and Smith formula,

$$\% \text{ ionic character} = 16(X_A - X_B) + 3.5(X_A - X_B)^2$$

④ Nomenclature of inorganic compounds

→ Prefix = less EN element — Cl_2O ✓
Suffix = More EN element. — OCl_2 X.

⑤ Bond strength ∝ Δ EN ∝ Bond polarity.

Eg - HF > HCl > HBr > HI.

∴ acidic strength - HI > HBr > HCl > HF.

8. X_m = Mulliken's EN X_p = Pauli's EN.

$$X_m = \frac{IP + EA}{2}$$

$$X_p = \frac{X_m}{2.8}$$

→ If IP, EA are in eV ⇒ $X_p = \frac{IP + EA}{5.6}$

→ If IP, EA are in Kcal/mol ⇒ $X_p = \frac{IP + EA}{2 \times 62.5}$