Periodic Classification of Elements

Definition:

Periodic classification is the arrangement of elements in such a way that elements with similar properties reappear at regular intervals.

Historical Development in Classification

Dobereiner's Triads (1817):

- Elements were grouped into triads (sets of 3).
- The atomic mass of the middle element was nearly the average of the other two.
  Example:
  Li (7), Na (23), K (39) → Mean of Li and K = (7+39)/2 = 23 ≈ Na

Newlands' Law of Octaves (1866):

- When elements are arranged in increasing order of atomic masses, every 8th element resembles the first.
- It worked well for lighter elements.

Lothar Meyer's Arrangement:

- Plotted atomic volume vs atomic mass.
- Showed periodic properties are functions of atomic volume.

Mendeleev's Periodic Table (1869):

- Based on atomic mass.

- Mendeleev's Periodic Law: *"Properties of elements are periodic functions of their atomic masses."*

- 8 groups and 7 periods, 63 elements arranged.

Defects:

- No proper position for hydrogen.

- Disrupted order in some elements (e.g., Ar-K).

- No place for isotopes.

- Grouping sometimes separated similar elements or combined dissimilar ones.

Modern Periodic Table (Moseley, 1913)

- Based on **atomic number** (Z), not atomic mass.

- **Modern Periodic Law**: *"Properties of elements are periodic functions of their atomic numbers."*

Key Features:

- 18 vertical groups and 7 horizontal periods.

- Elements grouped by outer electronic configuration.

- Period number = highest principal quantum number (n).

- Lanthanides and actinides placed separately.

Block Classification

- **s-block:** Groups 1 and 2 – Alkali and alkaline earth metals.
- **p-block:** Groups 13 to 18 – Includes non-metals, metalloids, and some metals.
- **d-block:** Groups 3 to 12 – Transition elements.
- **f-block:** Lanthanides (Z=58–71), Actinides (Z=90–103) – Inner transition elements.

Periodic Properties

1. Atomic Radius

- Distance from nucleus to outermost shell.
- Increases down a group (more shells).
- Decreases across a period (higher nuclear charge).
- Cation radius < atom < anion.
- Isoelectronic species: greater Z → smaller radius.

2. Ionisation Potential (IP)

- Energy required to remove an electron.
- Increases across a period, decreases down a group.
- IP order: s > p > d > f orbitals.
- Group 2 > Group 13; Group 15 > Group 16 due to stability.

3. Electron Affinity (EA)

- Energy released when an atom gains an electron.
- Increases across a period, decreases down a group.
- Cl > F > Br > I (F is small, repulsion reduces EA).

4. Electronegativity

- Tendency to attract electrons in a bond.
- Decreases down a group, increases across a period.
  F > O > Cl ≈ N > Br > C ≈ I > H

Group-Wise Properties

Group IA – Alkali Metals (Li, Na, K, Rb, Cs, Fr):

- Electronic configuration: ns^1
- Very reactive, soft metals, low MP/BP.
- Stored in paraffin; strong reducing agents.
- Form monovalent cations; ionic compounds.
- Flame test: Li-red, Na-yellow, K-pale violet, etc.
- Reactivity: Li < Na < K < Rb < Cs
- Li forms Li_2O, Na forms Na_2O_2, others form superoxides.

Group IIA – Alkaline Earth Metals (Be, Mg, Ca, Sr, Ba, Ra):

- Electronic configuration: ns^2
- Be and Mg don't show flame color (high IP).

- Flame colors: Ca–brick red, Sr–crimson, Ba–apple green.
- Hydroxide solubility increases down the group.
- Carbonates are less soluble than IA group.

Group IIIA (B, Al, Ga, In, Tl):

- ns^2np^1; Oxidation states: +3 and +1 (inert pair effect).
- Boron is metalloid; others are metals.
- Boron → glass industry; Al → thermite process.
- Al is amphoteric.
- Order of oxide basicity: B < Al < Ga < In < Tl

Group IVA (C, Si, Ge, Sn, Pb):

- ns^2np^2; Oxidation states: +4 and +2 (inert pair effect).
- C forms strong covalent bonds, catenation.
- CO_2 is linear, SiO_2 is a 3D network.
- Hydride stability: $CH_4 > SiH_4 > GeH_4 > SnH_4 > PbH_4$

Group VA (N, P, As, Sb, Bi):

- ns^2np^3; half-filled stability.
- Oxidation states: -3, +3, +5
- N shows anomalous properties (no d-orbital).
- N_2: triple bond; hydrides: $NH_3 > PH_3 > AsH_3 > SbH_3$ (boiling pt)

- N forms multiple oxides (NO, N_2O, NO_2, etc.)

Group VIA – Chalcogens (O, S, Se, Te, Po):

- ns^2np^4; Oxidation states: -2, +2, +4, +6
- O_2 is paramagnetic; ozone is an allotrope.
- Acidic character: $H_2O < H_2S < H_2Se < H_2Te$
- Volatility: $H_2O > H_2Te > H_2Se > H_2S$

Group VIIA – Halogens (F, Cl, Br, I, At):

- ns^2np^5; show -1 to +7 oxidation states.
- High reactivity and electron affinity.
- Oxidizing power: $F_2 > Cl_2 > Br_2 > I_2$
- HF is least acidic but most stable.

Zero Group – Noble Gases (He, Ne, Ar, Kr, Xe, Rn):

- ns^2np^6 (He is $1s^2$)
- Inert, monoatomic gases; Xe forms compounds ($XePtF_6$)
- Uses: He in balloons, Ne in lamps, Ar in bulbs, Rn in cancer therapy.

Transition Elements (d-block)

- Small atomic size, high melting points, variable valency.
- Colored compounds due to d–d transitions.

- Good conductors of heat and electricity.
- Used as catalysts.
- Mn: max oxidation state +7 in $KMnO_4$.
- Ag halides + sodium thiosulfate → used in photography.

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