Class 11 Chemistry

Chapter 3 – Classification of Elements and Periodicity in Properties

1. Introduction

The periodic classification of elements forms the basis of modern chemistry. When elements are arranged in order of increasing atomic number, the properties of elements repeat at regular intervals. This repetition is called periodicity. Cause of periodicity: Similar valence-shell electronic configurations appear after definite intervals. Modern Periodic Law (Moseley, 1913): "The physical and chemical properties of the elements are the periodic functions of their atomic numbers."

2. Historical Development of the Periodic Table

Dobereiner (1829): Law of Triads – three elements of similar properties arranged so that the atomic mass of the middle element is approximately the mean of the other two. Only a few triads known (e.g., Li-Na-K). Newlands (1866): Law of Octaves – when elements are arranged in increasing atomic mass, every eighth element has properties similar to the first. Valid only up to calcium. Mendeleev (1869): Periodic Law (based on atomic mass). Left gaps for undiscovered elements and predicted their properties. Great success but had anomalies (isotopes, Co–Ni, H position). Moseley (1913): X-ray studies showed atomic number, not atomic mass, is the true basis of periodicity. Led to the Modern Periodic Law.

3. Modern Long-Form Periodic Table

Basis: Increasing atomic number. Structure: 7 periods (horizontal rows), 18 groups (vertical columns), 118 known elements. Blocks according to valence-shell configuration: s-block (Groups 1–2), p-block (Groups 13–18), d-block (Groups 3–12), f-block (Lanthanides and Actinides).

4. Relationship between Electronic Configuration and Position

Period number corresponds to the principal quantum number (n) of the outermost shell. Group number depends on the number of valence electrons. Elements with similar outer-shell configurations lie in the same group and exhibit similar chemical properties.

5. Important Periodic Properties and Their Trends

5.1 Atomic Radius: Decreases across a period; increases down a group. 5.2 Ionic Radius: Cations smaller than parent atoms; anions larger. Isoelectronic ions shrink with increasing nuclear charge. 5.3 Ionization Enthalpy: Increases across a period, decreases down a group; exceptions Be>B, N>O. 5.4 Electron Gain Enthalpy: Becomes more negative across a period; less negative down a group; exceptions for noble gases. 5.5 Electronegativity: Increases across a period, decreases down a group; Fluorine most electronegative. 5.6 Metallic and Non-metallic Character: Metallic decreases across a period, increases down a group. 5.7 Valency: Increases 1→4 then decreases to 0 across a period.

6. Periodic Trends in Chemical Reactivity

s-Block: Reactivity increases down group. p-Block: Reactivity decreases down group. Transition metals: Variable valency, moderate reactivity. Noble gases: Inert.

7. Anomalous Behaviour of Second-Period Elements

Reasons: Small size, high electronegativity, high ionization enthalpy, absence of d-orbitals. Examples: Li forms Li_2O; $BeCl_2$ covalent unlike $MgCl_2$ ionic.

8. Diagonal Relationship

Li–Mg and Be–Al show similarities due to similar charge/radius ratio (polarizing power).

9. Shielding Effect and Effective Nuclear Charge

Shielding effect: Inner electrons reduce nuclear attraction; s>p>d>f. Effective nuclear charge (Z_{eff}) = Z – S (S = screening constant).

10. Periodic Trends – Summary Table

Across a Period: Atomic radius ↓, Ionization enthalpy ↑, Electron gain enthalpy more negative, Electronegativity ↑, Metallic character ↓. Down a Group: Atomic radius ↑, Ionization enthalpy ↓, Electron gain enthalpy less negative, Electronegativity ↓, Metallic character ↑.

11. Anomalies and Exceptions

Be>B, N>O in ionization energy; Fluorine less negative electron gain enthalpy than Cl; Hydrogen behaves uniquely.

12. Types of Elements Based on Electronic Configuration

s-block ($ns^{1–2}$): Highly reactive metals. p-block ($ns^2 np^{1–6}$): Metals, nonmetals, metalloids. d-block ($(n–1)d^{1–10} ns^{1–2}$): Transition metals. f-block ($(n–2)f^{1–14} (n–1)d^{0–1} ns^2$): Inner transition elements.

13. Applications of the Periodic Table

Predicting properties, explaining bonding, valency, and location of new elements.

14. Common Examination Questions

Very Short Answer: Define Modern Periodic Law, shielding effect, etc. Short Answer: Explain radius trend, Be>B anomaly. Long Answer: Describe periodic trends, explain table structure.

15. Key Points for Rapid Revision

Periodicity due to valence configuration; atomic number governs order; across period nonmetallic character rises; down group metallic character increases.

16. Formulae and Quantitative Relationships

$Z_{eff} = Z – S$; $IE_1 < IE_2 < IE_3$; Atomic radius $\propto 1/Z_{eff}$; $\Delta\chi > 1.7 \rightarrow$ Ionic bond.

17. Summary

The periodic table arranges all known elements by atomic number; periodic trends determine chemical behaviour and bonding characteristics.