

Chapter 3: Classification of Elements and Periodicity in Properties

Comprehensive Study Notes

Class 11 Chemistry - NCERT Based

EXAM SPRINT - Complete Coverage for NEET and JEE Examinations

Introduction

The Periodic Table is arguably the most important concept in chemistry, both in principle and in practice. It provides a succinct organization of the whole of chemistry and demonstrates that chemical elements are not random entities but display trends and lie together in families. Understanding the Periodic Table is essential for anyone who wishes to understand how the world is built from fundamental building blocks - the chemical elements.

"The properties of the elements are a periodic function of their atomic numbers." - Modern Periodic Law

3.1 Why Do We Need to Classify Elements?

Historical Context:

- **1800:** Only 31 elements known
- **1865:** 63 elements identified
- **Present:** 118 elements known (recent ones are man-made)

Need for Classification:

1. **Systematic Study:** Too difficult to study chemistry of all elements individually

2. **Pattern Recognition:** Rationalize known chemical facts
3. **Prediction:** Predict properties of new elements
4. **Organization:** Ease understanding and learning

3.2 Genesis of Periodic Classification

Historical Development Timeline:

3.2.1 Dobereiner's Triads (1829)

Key Concept: Groups of three elements with similar properties

Dobereiner's Triads Table:

Element	Atomic Weight	Element	Atomic Weight	Element	Atomic Weight
Li	7	Ca	40	Cl	35.5
Na	23	Sr	88	Br	80
K	39	Ba	137	I	127

Observations:

- Middle element's atomic weight $\approx$ average of other two
- Properties of middle element intermediate between other two
- **Law of Triads:** Limited application, worked only for few elements

3.2.2 Newlands' Law of Octaves (1865)

Key Concept: Every eighth element had similar properties

Newlands' Octaves Table:

Element	Li	Be	B	C	N	O	F
At. wt.	7	9	11	12	14	16	19
Element	Na	Mg	Al	Si	P	S	Cl
At. wt.	23	24	27	29	31	32	35.5

Limitations:

- Valid only up to calcium
- No provision for noble gases
- Some dissimilar elements grouped together

3.2.3 Mendeleev's Periodic Law (1869)

Key Scientist: Dmitri Mendeleev (Russian chemist)

Mendeleev's Periodic Law: *"The properties of the elements are a periodic function of their atomic weights."*

Revolutionary Features:

1. **Gaps for Undiscovered Elements:** Left spaces for unknown elements
2. **Property-Based Arrangement:** Ignored atomic weight order when properties demanded
3. **Predictions:** Described properties of undiscovered elements

Famous Predictions:

Property	Eka-aluminium (Predicted)	Gallium (Found)	Eka-silicon (Predicted)	Germanium (Found)
Atomic weight	68	70	72	72.6
Density (g/cm ³)	5.9	5.94	5.5	5.36
Melting point	Low	302.93 K	High	1231 K
Oxide formula	E ₂ O ₃	Ga ₂ O ₃	EO ₂	GeO ₂

3.3 Modern Periodic Law and Present Form of Periodic Table

3.3.1 Moseley's Contribution (1913)

Discovery: Atomic number is more fundamental than atomic mass **Method:** X-ray spectroscopy showed linear relationship between $\sqrt{\nu}$ and atomic number

Modern Periodic Law:

"The physical and chemical properties of the elements are periodic functions of their atomic numbers."

Significance:

- Atomic number = Nuclear charge = Number of protons = Number of electrons (in neutral atom)
- Electronic configuration determines properties
- Explains periodicity scientifically

3.3.2 Modern Periodic Table Structure

Key Features:

- **7 Periods** (horizontal rows)

- **18 Groups** (vertical columns)
- **IUPAC Numbering:** Groups 1-18

Period-Electron Relationship:

Period	n value	Maximum Electrons	Number of Elements
1	1	2	2
2	2	8	8
3	3	8	8
4	4	18	18
5	5	18	18
6	6	32	32
7	7	32	32 (incomplete)

3.4 Nomenclature of Elements with Atomic Numbers > 100

IUPAC Systematic Nomenclature:

Principle: Name derived from atomic number using numerical roots

Numerical Roots Table:

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

Example: Element 120 → un-bi-nil-ium → Unbinilium (Ubn)

3.5 Electronic Configurations and Types of Elements

3.5.1 Block Classification

Elements classified based on subshell being filled:

s-Block Elements (Groups 1-2)

Electronic Configuration: ns^1 (Group 1), ns^2 (Group 2) **Properties:**

- Highly reactive metals
- Low ionization enthalpies
- Form ionic compounds (except Li, Be)
- Metallic character increases down group

p-Block Elements (Groups 13-18)

Electronic Configuration: ns^2np^1 to ns^2np^6 **Properties:**

- Representative elements
- Include metals, non-metals, metalloids
- Non-metallic character increases across period
- Noble gases have complete octets

d-Block Elements (Groups 3-12)

Electronic Configuration: $(n-1)d^{1-10}ns^{0-2}$ **Properties:**

- Transition elements
- Form colored compounds
- Variable oxidation states
- Paramagnetic properties
- Catalytic activity

f-Block Elements

Electronic Configuration: $(n-2)f^{1-14}(n-1)d^{0-1}ns^2$ **Series:**

- **Lanthanoids:** Ce (58) to Lu (71)
- **Actinoids:** Th (90) to Lr (103) **Properties:**
- Inner transition elements
- Similar properties within series
- Radioactive (actinoids)

3.5.2 Metals, Non-metals, and Metalloids

Distribution:

- **Metals:** >78% of known elements (left side of periodic table)
- **Non-metals:** <20 elements (top right)
- **Metalloids:** Along zig-zag line (Si, Ge, As, Sb, Te)

Property Trends:

- Metallic character: Increases down group, decreases across period
- Non-metallic character: Decreases down group, increases across period

3.6 Periodic Trends in Properties

3.6.1 Atomic Radius

Definition: Half the internuclear distance between identical atoms

Types:

- **Covalent radius:** Half of bond length in homonuclear molecule
- **Metallic radius:** Half of internuclear distance in metallic crystal

Trends:

1. **Across period:** Decreases (Li: 152 pm → F: 64 pm)
2. **Down group:** Increases (Li: 152 pm → Cs: 262 pm)

Explanation:

- Across period: Effective nuclear charge increases
- Down group: Addition of electron shells

3.6.2 Ionic Radius

Key Concepts:

- **Cation:** Smaller than parent atom (electron loss)
- **Anion:** Larger than parent atom (electron gain)
- **Isoelectronic species:** Same electron number, different sizes

Examples:

- Na (186 pm) $\rightarrow$ Na⁺ (95 pm)
- F (64 pm) $\rightarrow$ F⁻ (136 pm)
- Isoelectronic: O²⁻ > F⁻ > Na⁺ > Mg²⁺ (all have 10 electrons)

3.6.3 Ionization Enthalpy (IE)

Definition: Energy required to remove electron from gaseous atom

Mathematical Expression: $X(g) \rightarrow X^+(g) + e^- \Delta_i H_1$ (First IE) $X^+(g) \rightarrow X^{2+}(g) + e^- \Delta_i H_2$ (Second IE)

Trends:

1. **Across period:** Generally increases
2. **Down group:** Decreases
3. **Successive IEs:** Always increase ($IE_1 < IE_2 < IE_3$)

Anomalies:

- Be > B (2s vs 2p electron removal)
- N > O (electron-electron repulsion in O)

3.6.4 Electron Gain Enthalpy (EGE)

Definition: Energy change when electron is added to gaseous atom

Mathematical Expression: $X(g) + e^- \rightarrow X^-(g) \Delta_e gH$

Trends:

1. **Across period:** Becomes more negative
2. **Down group:** Becomes less negative

Special Cases:

- Halogens: Most negative EGE values
- Noble gases: Positive EGE values
- F vs Cl: Cl has more negative EGE than F (electron-electron repulsion)

3.6.5 Electronegativity

Definition: Ability of atom to attract shared electrons

Pauling Scale: F = 4.0 (highest), Cs = 0.7 (lowest)

Trends:

1. **Across period:** Increases
2. **Down group:** Decreases

Relationship with Properties:

- Directly related to non-metallic character
- Inversely related to metallic character

3.7 Periodic Trends in Chemical Properties

3.7.1 Valency and Oxidation States

Representative Elements:

- Valency = Number of valence electrons OR (8 - valence electrons)
- Examples: Na (1), Mg (2), Al (3), C (4), N (3,5), O (2,6), F (1)

3.7.2 Chemical Reactivity Patterns

Reactivity Trends:

1. **Across period:** High at extremes (Groups 1 and 17), low in center
2. **Group trends:**
 - Alkali metals: Reactivity increases down group
 - Halogens: Reactivity decreases down group

Oxide Character:

- **Basic oxides:** Left side ($\text{Na}_2\text{O} + \text{H}_2\text{O} \rightarrow 2\text{NaOH}$)
- **Acidic oxides:** Right side ($\text{Cl}_2\text{O}_7 + \text{H}_2\text{O} \rightarrow 2\text{HClO}_4$)
- **Amphoteric oxides:** Center (Al_2O_3)
- **Neutral oxides:** Transition region (CO, NO)

3.7.3 Anomalous Properties of Second Period

Elements: Li to F

Reasons for Anomalous Behavior:

1. **Small size**
2. **High charge/radius ratio**

3. **High electronegativity**
4. **Limited valence orbitals** (only 2s, 2p available)
5. **Ability to form π - π multiple bonds**

Examples:

- Li behaves more like Mg than Na
- Be behaves more like Al than Mg
- BF_4^- (max covalency 4) vs AlF_6^{3-} (covalency 6)

NEET & JEE Specific Important Points

High-Yield Topics for NEET:

1. **Electronic Configuration Rules**
2. **Periodic Trends** (atomic radius, IE, EGE, electronegativity)
3. **Block Classification**
4. **Anomalous Properties**
5. **Chemical Reactivity Patterns**

High-Yield Topics for JEE:

1. **Mathematical Calculations** (IE, EGE problems)
2. **Shielding Effect and Effective Nuclear Charge**
3. **Diagonal Relationships**
4. **Advanced Trends** (lanthanoid contraction, inert pair effect)
5. **Chemical Bonding Correlations**

Critical Numerical Concepts:

1. **Effective Nuclear Charge:** $Z_{\text{eff}} = Z - \sigma$ (σ = shielding constant)
2. **Slater's Rules** for calculating shielding
3. **Size Relationships:**
 - Cation < Atom < Anion
 - For isoelectronic species: Size $\propto$ 1/Nuclear charge

Memory Aids and Mnemonics

Electronic Configuration Order:

"1s 2s 2p 3s 3p 4s 3d 4p 5s 4d 5p 6s 4f 5d 6p 7s 5f 6d 7p"

Block Elements:

- **s-block:** "Soft metals" (Groups 1-2)
- **p-block:** "Periodic properties" (Groups 13-18)
- **d-block:** "Dense transition metals" (Groups 3-12)
- **f-block:** "Rare earth elements"

Periodic Trends Memory:

"As you go RIGHT and UP, atoms get TIGHT"

- Right → Across period → Size decreases, IE increases
- Up → Up the group → Size decreases, IE increases

Common NEET/JEE Question Patterns

1. Electronic Configuration Questions:

- Write configurations for elements/ions
- Identify block, period, group from configuration
- Find number of unpaired electrons

2. Periodic Trend Problems:

- Compare atomic/ionic radii
- Arrange elements by IE/EGE/electronegativity
- Explain anomalies in trends

3. Chemical Property Correlations:

- Relate electronic configuration to chemical behavior
- Predict compound formulas based on periodic trends
- Explain reactivity patterns

4. Numerical Problems:

- Calculate effective nuclear charge
- Energy transition calculations
- Size comparisons using mathematical relationships

Practice Questions for Competitive Exams

Multiple Choice Questions:

1. Which electronic configuration represents the most stable ion? a) $1s^2 2s^2 2p^6 3s^1$ b) $1s^2 2s^2 2p^6$ c) $1s^2 2s^2 2p^5$ d) $1s^2 2s^2 2p^6 3s^2 3p^1$

2. The correct order of atomic radii is: a) $\text{Na} > \text{Mg} > \text{Al} > \text{Si}$ b) $\text{Si} > \text{Al} > \text{Mg} > \text{Na}$ c) $\text{Al} > \text{Mg} > \text{Na} > \text{Si}$ d) $\text{Mg} > \text{Na} > \text{Al} > \text{Si}$

3. Which element has the highest first ionization enthalpy? a) N b) O c) F d) Ne

Short Answer Questions:

1. Explain why the second ionization enthalpy of sodium is much higher than the first.
2. Why do transition elements show variable oxidation states?
3. Compare the chemical reactivity of alkali metals and halogens.

Long Answer Questions:

1. Derive the periodic trends in atomic radii and ionization enthalpies across a period and down a group.
2. Explain the anomalous properties of second-period elements with suitable examples.
3. Describe the development of the periodic table from Dobereiner to the modern periodic law.

Summary Tables

Periodic Trends Summary:

Property	Across Period	Down Group	Reason
Atomic Radius	Decreases	Increases	Nuclear charge vs electron shells
Ionization Enthalpy	Increases	Decreases	Nuclear attraction vs shielding
Electron Gain Enthalpy	More negative	Less negative	Nuclear attraction vs size
Electronegativity	Increases	Decreases	Similar to ionization enthalpy
Metallic Character	Decreases	Increases	Electron loss tendency

Block Classification Summary:

Block	Groups	Elements	Key Properties
s	1-2	H, Li-Fr, Be-Ra	Reactive metals, ionic compounds
p	13-18	B-Rn	Metals, non-metals, metalloids
d	3-12	Sc-Hg	Transition metals, colored ions
f	Lanthanoids/Actinoids	Ce-Lu, Th-Lr	Inner transition, similar properties

Advanced Topics for JEE Advanced

1. Lanthanoid Contraction:

- **Definition:** Steady decrease in size across lanthanoid series
- **Cause:** Poor shielding by f-electrons
- **Consequences:** Similar sizes of 4d and 5d elements

2. Inert Pair Effect:

- **Definition:** Reluctance of ns^2 electrons to participate in bonding
- **Occurs in:** Heavier p-block elements (Pb^{2+} vs Pb^{4+})
- **Increases down:** Group 13-15

3. Diagonal Relationships:

- **Li-Mg:** Similar polarizing power
- **Be-Al:** Similar charge/size ratio
- **B-Si:** Similar electronegativity

EXAM SPRINT - Master Classification and Periodicity

Key Success Strategy: Focus on understanding the logical development of periodic classification and the underlying electronic structure basis. Master periodic trends with their explanations rather than just memorizing. Practice numerical problems involving effective nuclear charge and size calculations.

Essential for Competitive Exams:

1. Electronic configuration mastery
2. Trend explanation ability
3. Anomaly identification skills
4. Chemical property predictions
5. Mathematical problem-solving

Time Management Tips:

- Trends questions: 1-2 minutes each
- Electronic configuration: 30 seconds
- Chemical property predictions: 2-3 minutes
- Numerical problems: 3-4 minutes

Source: NCERT Chemistry Class 11, Chapter 3 - Enhanced with competitive exam focus and advanced concepts for NEET and JEE preparation.