EDUVISTA

CLASS 11 CHEMISTRY

UNIT 3
CLASSIFICATION OF ELEMENTS AND PERIODICITY IN PROPERTIES

EDUVISTA-STYLE SMART NOTES

Concepts • Trends • Exceptions • Exam Focus • NCERT Practice

Based on the uploaded NCERT PDF — Unit 3, Reprint 2026–27

Learn → Understand → Practice → Test → Improve

1. WHY CLASSIFY ELEMENTS?

The chapter explains that as the number of known elements increased, studying each element and its innumerable compounds individually became difficult. Classification provides a systematic organisation, rationalises known facts and can help predict new ones.

2. GENESIS OF PERIODIC CLASSIFICATION

Scientist	Contribution	Key limitation / value
Döbereiner	Triads of three similar elements; middle atomic weight was approximately midway.	Worked only for a few groups.
de Chancourtois	Arranged elements by increasing atomic weight on a cylindrical arrangement.	Early display of periodicity.
Newlands	Law of Octaves: every eighth element showed similarity.	Mainly worked up to calcium.
Lothar Meyer	Studied recurring physical-property patterns and developed a table close to the modern form.	Independent contribution to periodic classification.
Mendeleev	Used atomic weight and chemical properties; placed similar elements together and left gaps for undiscovered elements.	Successful predictions of gallium and germanium.

Mendeleev’s Periodic Law

The properties of the elements are a periodic function of their atomic weights.

Mendeleev sometimes ignored strict atomic-weight order to keep chemically similar elements together. He also left gaps and predicted properties of undiscovered elements.

3. MODERN PERIODIC LAW

Henry Moseley showed that atomic number is more fundamental than atomic mass. The Modern Periodic Law therefore states:

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

Modern long form

• Horizontal rows = periods; vertical columns = groups/families.

• IUPAC groups are numbered 1–18.

• Seven periods; lengths are 2, 8, 8, 18, 18, 32 and an incomplete 7th period with theoretical maximum 32.

• Period number corresponds to the highest principal quantum number n.

• Lanthanoids and actinoids are shown separately at the bottom.

• Same-group elements have similar outer electronic configurations and similar chemical properties.

4. IUPAC NOMENCLATURE FOR Z > 100

Digit	Root	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

PDF example: Z = 120 → un + bi + nil + ium = **unbinilium**; symbol **Ubn**.

5. ELECTRONIC CONFIGURATION & PERIODS

Electronic configuration is the distribution of electrons into orbitals. The period indicates the principal quantum number of the outermost/valence shell.

Period	Max filling described in PDF	No. of elements
1	1s	2
2	2s, 2p	8
3	3s, 3p	8
4	4s, 3d, 4p	18
5	5s, 4d, 5p	18
6	6s, 4f, 5d, 6p	32
7	7s, 5f, 6d, 7p	Incomplete; theoretical max 32

Why 18 elements in the 5th period?

For n = 5, the relevant orbitals include 5s, 4d and 5p. Total available orbitals = 1 + 5 + 3 = 9; each orbital holds 2 electrons → 18 electrons → 18 elements.

6. BLOCKS

Block	General outer configuration	Location
s	$ns^{1–2}$	Groups 1–2
p	$ns^2np^{1–■}$	Groups 13–18
d	$(n–1)d^{1–1■}$ $ns^{■–2}$	Groups 3–12
f	$(n–2)f^{1–1■}$ $(n–1)d^{■–1}$ ns^2	Lanthanoids & actinoids

Special cases: He has $1s^2$ but is positioned in Group 18 because of its filled valence shell. Hydrogen is kept separately because it can show Group 1-like and Group 17-like behaviour.

7. s, p, d, f BLOCKS & TYPES OF ELEMENTS

s-block

• Groups 1 and 2; ns^1 and ns^2 configurations.

• Reactive metals with low ionization enthalpies.

• Form 1+ ions in Group 1 and 2+ ions in Group 2.

• Metallic character and reactivity generally increase down the group.

• Except for Li and Be, compounds are predominantly ionic.

p-block

• Groups 13–18; ns^2np^1 to ns^2np■.

• Includes metals, non-metals and metalloids.

• Group 17 = halogens; Group 16 = chalcogens.

• Group 18 noble gases have filled valence shells and very low reactivity.

d-block

• Groups 3–12; filling of inner d orbitals.

• These are transition elements.

• Variable oxidation states are characteristic.

f-block

• 4f series = lanthanoids; 5f series = actinoids.

• Placed separately at the bottom of the table to preserve its structure.

• Actinoids include many man-made radioactive elements.

Metals • Non-metals • Metalloids

• Metals mainly occupy the left; they are generally good conductors, malleable and ductile.

• Non-metals are mainly toward the top-right; they are generally poor conductors and many are brittle.

• Metalloids lie around the zig-zag boundary: Si, Ge, As, Sb, Te.

8. PERIODIC TRENDS — MASTER MAP

Property	Across a period	Down a group
Atomic radius	↓	↑
Ionic radius	Generally follows atomic-size trend; compare charge/electrons carefully	Generally ↑
Ionization enthalpy	↑ (generally)	↓
Electron gain enthalpy	Generally more negative	Generally less negative
Electronegativity	↑	↓
Metallic character	↓	↑
Non-metallic character	↑	↓

Atomic radius

• Decreases across a period because effective nuclear charge increases while electrons enter the same valence shell.

• Increases down a group because new energy levels are added and shielding increases.

Ionic radius

• Cation is smaller than its parent atom because it has fewer electrons with the same nuclear charge.

• Anion is larger than its parent atom because added electrons increase electron–electron repulsion.

• Isoelectronic species have the same number of electrons; greater nuclear charge gives smaller radius.

Ionization enthalpy

Energy required to remove an electron from an isolated gaseous atom in its ground state. It is positive, and successive ionization enthalpies increase.

Key exceptions

• Be > B because B loses a 2p electron whereas Be loses a more strongly held 2s electron.

• N > O because O has pairing in a 2p orbital, increasing electron–electron repulsion.

9. ELECTRON GAIN ENTHALPY • EN • VALENCE

Electron Gain Enthalpy

• Energy change when an electron is added to an isolated gaseous atom.

• Across a period it generally becomes more negative.

• Noble gases have rather positive electron gain enthalpy because of their filled shells.

• Second electron gain enthalpy is less favourable because an electron is added to an already negative ion.

Electronegativity

• Tendency of an atom in a molecule to attract the shared electron pair.

• Generally increases left → right and decreases down a group.

• Directly related to non-metallic character and inversely related to metallic character.

• It is not constant for an element; it depends on the atom to which it is bonded.

Valence / oxidation state

• For representative elements, valence is usually the number of outermost electrons or 8 minus that number.

• Oxidation state is assigned using electronegativity considerations.

• PDF examples: SiBr■ and Al■S■.

10. CHEMICAL REACTIVITY

• Within a period, reactivity is high at the two extremes and lowest near the centre.

• Group 1 metals react mainly by electron loss; Group 17 non-metals mainly by electron gain.

• Highly reactive elements generally occur in combined form in nature.

• Metallic character decreases left → right; non-metallic character increases.

Oxides

• Left extreme → basic oxides, e.g. Na■O.

• Right extreme → acidic oxides, e.g. Cl■O■.

• Centre → amphoteric or neutral oxides, e.g. Al■O■, As■O■, CO, NO, N■O.

• Na■O + H■O → 2NaOH; Cl■O■ + H■O → 2HClO■.

11. ANOMALOUS BEHAVIOUR & DIAGONAL RELATIONSHIP

The first element of several s- and p-block groups differs from later members. The PDF attributes this to small size, large charge/radius ratio, high electronegativity and limited valence orbitals.

Why the first member behaves differently

• Small atomic size.

• Large charge/radius ratio.

• High electronegativity.

• Only four valence orbitals (2s, 2p) are available in the second period.

• Maximum covalency of the first member is 4.

• Second-period p-block elements have a greater ability to form pπ–pπ multiple bonds.

Diagonal relationship

The chapter describes similarities between certain first members and the second element of the next group, associated with comparable size and charge-density effects. Lithium–magnesium and beryllium–aluminium are important examples discussed in the chapter.

12. QUICK COMPARISON BOX

Question	Answer
Largest / smallest among Mg, Mg^{2}■, Al, Al^{3}■	Largest Mg; smallest Al^{3}■
Increasing metallic character: P, Si, Be, Mg, Na	P < Si < Be < Mg < Na
Group 1 reactivity down group	Li < Na < K < Rb < Cs
Group 17 reactivity down group	F > Cl > Br > I
Period number represents	Highest principal quantum number n
Same group similarity	Similar valence-shell electronic configuration

13. COMMON EXAM MISTAKES

• Do not replace atomic number with atomic mass in the Modern Periodic Law.

• Do not treat ionization enthalpy as negative; removal of an electron requires energy.

• Do not blindly apply trends without checking Be/B and N/O exceptions.

• For isoelectronic species, compare nuclear charge to rank radii.

• Remember that hydrogen is a special case and helium is positioned with Group 18.

14. NCERT PRACTICE — MUST DO

The uploaded PDF includes a full exercise set. Prioritise these for active recall and test preparation:

• 3.1–3.3: organisation of the Periodic Table and Mendeleev vs Modern Periodic Law.

• 3.4–3.6: quantum-number justification, locating elements and atomic numbers.

• 3.8–3.13: group similarity, atomic/ionic radii and isoelectronic species.

• 3.14–3.16: definitions and ionization-enthalpy calculations/exceptions.

• 3.17–3.25: ionization enthalpy, electron gain enthalpy and electronegativity.

• 3.26–3.30: metals/non-metals, reactivity, block configurations and locating elements.

• 3.31–3.34: data-based identification, compound formulas and modern-table MCQs.

• 3.35–3.40: higher-order questions on valence shell, isoelectronic size, ionization, metallic/non-metallic character and reactivity.

15. ONE-PAGE REVISION

Across: radius ↓ | IE ↑ | EGE more negative | EN ↑ | metallic ↓ | non-metallic ↑
Down: radius ↑ | IE ↓ | EGE generally less negative | EN ↓ | metallic ↑ | non-metallic ↓
Reactivity: Group 1 increases down the group; Group 17 decreases down the group; within a period, reactivity is high at the extremes and low near the centre.

EDUVISTA
Learn • Practice • Succeed

Prepared from the uploaded NCERT Unit 3 PDF. Use the original PDF for exact diagrams, full tables and complete worked examples.