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FSc / ICS Part-II Chemistry — Punjab Board

CHAPTER 1

Periodic Classification of Elements and Periodicity

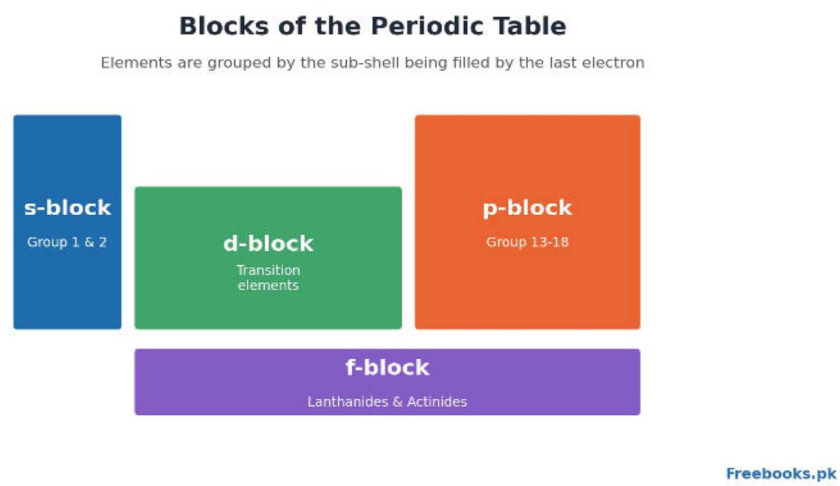

Figure 1.1 — The four blocks (s, p, d, f) of the periodic table.

1. Chapter Overview

The periodic table is one of the most powerful tools in chemistry. It arranges all known elements in a simple, organised pattern so that elements with similar properties fall together. In this chapter we study how this arrangement developed over time, how the modern periodic table is built, and how the properties of elements change (repeat in a regular way) as we move across a period or down a group. This regular repetition of properties is called periodicity.

Once you understand the pattern behind the periodic table, you no longer need to memorise the behaviour of each element separately. Instead, you can predict the size of an atom, how easily it loses or gains electrons, and whether its oxide is acidic or basic — simply from its position in the table. This makes Chapter 1 the foundation for the whole of Part-II Chemistry.

2. Learning Objectives

After studying this chapter, you should be able to:

- Describe the historical development of the periodic table from Dobereiner to Moseley.
- State the modern periodic law and explain why atomic number is the true basis of classification.
- Identify periods, groups, and the s-, p-, d- and f-blocks in the periodic table.
- Define and explain periodicity and the cause behind it.
- Explain the trends in atomic radius, ionic radius, ionization energy, electron affinity and electronegativity across a period and down a group.
- Relate shielding effect and effective nuclear charge to periodic trends.
- Predict the acidic or basic nature of oxides and hydrides of elements.
- Justify the anomalous position of hydrogen in the periodic table.

3. Key Concepts

3.1 Historical Background

Chemists always wanted to organise elements in a way that reflects their properties. Several attempts led to the modern periodic table:

Dobereiner's Triads (1829)

The German chemist Johann Dobereiner grouped elements into sets of three, called triads, having similar chemical properties. He noticed that the atomic mass of the middle element was roughly the average of the other two. For example, in the triad Lithium (7), Sodium (23) and Potassium (39), the average of 7 and 39 is 23 — the mass of sodium. The idea was useful but only a few elements could be arranged this way.

Newlands' Law of Octaves (1864)

The English chemist John Newlands arranged elements in order of increasing atomic mass and found that every eighth element repeated the properties of the first — like the notes of a musical scale. This was called the Law of Octaves. It worked for lighter elements but broke down for heavier ones, so it was not widely accepted at the time.

Mendeleev's Periodic Table (1869)

The Russian chemist Dmitri Mendeleev arranged the elements in order of increasing atomic mass in horizontal periods and vertical groups. His periodic law stated: 'The properties of elements are a periodic function of their atomic masses.' Mendeleev's great success was that he left gaps for undiscovered elements and even predicted their properties (such as eka-silicon, later found as germanium).

Defects of Mendeleev's Table

Position of hydrogen was not fixed (it resembles both Group IA and Group VIIA).

Isotopes of an element had no separate place, although they have different masses.

Some heavier elements were placed before lighter ones (anomalous pairs) to keep similar elements together, e.g. Argon (40) before Potassium (39).

Moseley's Work and the Modern Periodic Law (1913)

The English physicist Henry Moseley showed by X-ray studies that the **atomic number** (number of protons), not the atomic mass, is the fundamental property of an element. This corrected the anomalies in Mendeleev's table and gave us the **Modern Periodic Law:** 'The properties of elements are a periodic function of their atomic numbers.'

3.2 The Modern Periodic Table

In the modern (long form) periodic table, elements are arranged in order of increasing atomic number. The horizontal rows are called periods and the vertical columns are called groups.

Feature	Description
Periods	7 horizontal rows. Period number = number of the outermost (valence) shell.
Groups	18 vertical columns. Elements in a group have the same number of valence electrons and similar properties.
Representative elements	s-block and p-block elements (Groups IA–VIIA and Group 0).
Transition elements	d-block elements, placed in the middle of the table.
Inner transition	f-block elements: lanthanides and actinides, placed at the bottom.

Important Families (Groups)

Alkali metals: Group IA — Li, Na, K, Rb, Cs, Fr. Very reactive, soft metals.

Alkaline earth metals: Group IIA — Be, Mg, Ca, Sr, Ba, Ra. Reactive metals, but less than alkali metals.

Chalcogens: Group VIA — O, S, Se, Te, Po. The 'ore-forming' family.

Halogens: Group VIIA — F, Cl, Br, I, At. Very reactive non-metals.

Noble (inert) gases: Group 0 / VIIIA — He, Ne, Ar, Kr, Xe, Rn. Very unreactive due to complete outer shells.

3.3 Classification into Blocks

Depending on which sub-shell (orbital) receives the last electron, elements are divided into four blocks:

s-block: Last electron enters an s-orbital. Groups IA and IIA. Soft, reactive metals.

p-block: Last electron enters a p-orbital. Groups IIIA–VIIA and Group 0. Contains metals, non-metals and metalloids.

d-block: Last electron enters a d-orbital. The transition elements. Hard metals, form coloured compounds.

f-block: Last electron enters an f-orbital. Lanthanides and actinides (inner transition elements).

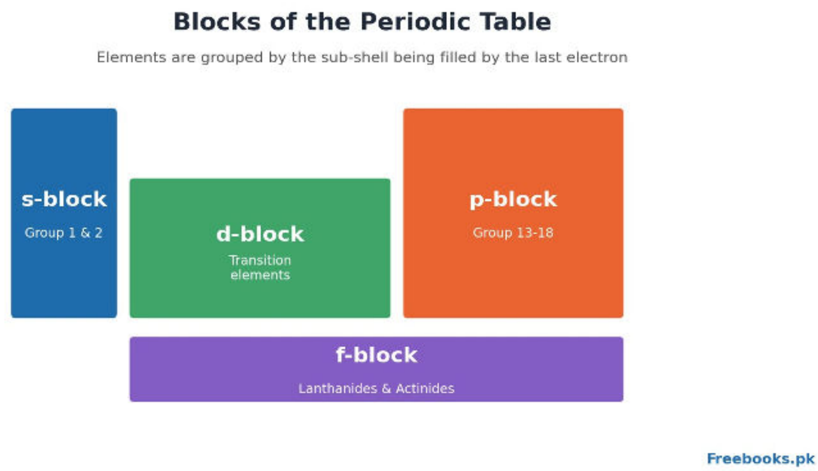

Figure 1.2 — Position of the s, p, d and f blocks.

3.4 Periodicity and Its Cause

Periodicity is the regular repetition of physical and chemical properties of elements after certain regular intervals when they are arranged by increasing atomic number. **Cause of periodicity:** it is due to the periodic repetition of the same outer (valence) electronic configuration. For example, all alkali metals have one electron in their outermost shell (ns^1), so they show similar properties.

3.5 Periodic Properties and Their Trends

Two ideas explain almost every trend in this chapter:

Effective nuclear charge (Zeff): the net positive charge actually felt by the valence electrons after the inner electrons partly cancel the nuclear pull. Higher Zeff pulls electrons in more strongly.

Shielding (screening) effect: the inner-shell electrons repel the outer electrons and reduce the nuclear attraction felt by them. More inner shells means more shielding.

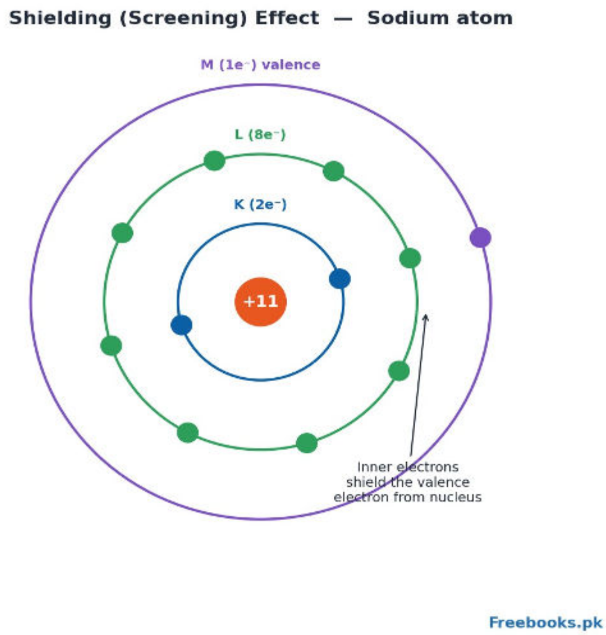

Figure 1.3 — Shielding of the valence electron in a sodium atom by inner shells.

Atomic Radius

The atomic radius is the distance from the nucleus to the outermost shell. Across a period it decreases because the nuclear charge increases while electrons are added to the same shell, pulling the shell inward. Down a group it increases because a new shell is added at each step.

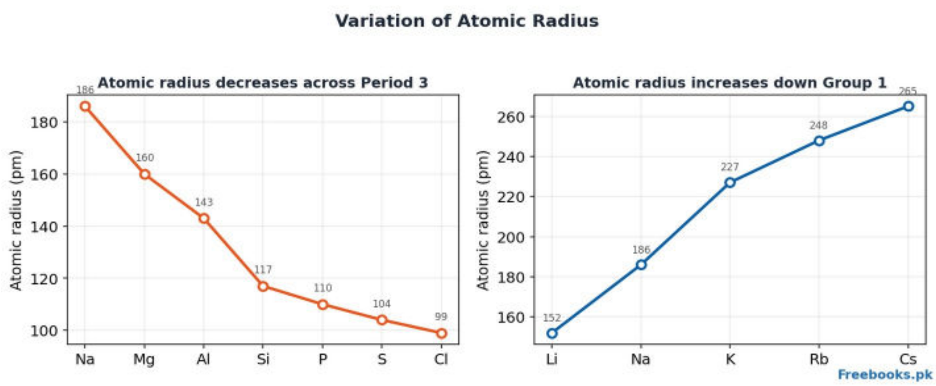

Figure 1.4 — Atomic radius decreases across a period and increases down a group.

Ionic Radius

A cation (positive ion) is always smaller than its parent atom because it has lost electrons and often a whole shell, so the remaining electrons are pulled in more tightly. An anion (negative ion) is always larger than its parent atom because added electrons increase electron–electron repulsion and the same nuclear charge now holds more electrons.

Ionization Energy

Ionization energy is the minimum energy required to remove the most loosely bound electron from an isolated gaseous atom. Across a period it increases (smaller size + higher Zeff hold electrons tightly). Down a group it decreases (larger size + more shielding make removal easier).

Small dips occur — for example, boron's value is lower than beryllium's because the 2p electron is easier to remove than a 2s electron.

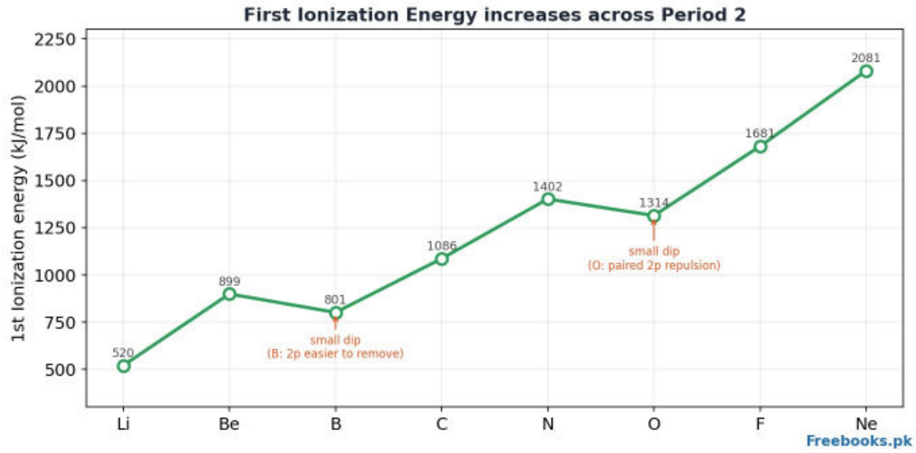

Figure 1.5 — First ionization energy across Period 2 (note the small dips at B and O).

Electron Affinity

Electron affinity is the energy released when an electron is added to an isolated gaseous atom. Generally it increases across a period (atoms more readily accept an electron) and decreases down a group. Halogens have the highest electron affinities, while noble gases have almost zero because their shells are already complete.

Electronegativity

Electronegativity is the tendency of an atom to attract the shared pair of electrons in a covalent bond towards itself. It increases across a period and decreases down a group. Fluorine is the most electronegative element.

Nature of Oxides and Hydrides

Across a period the character of oxides changes from basic (metals, e.g. Na_2O) through amphoteric (e.g. Al_2O_3) to acidic (non-metals, e.g. SO_3, Cl_2O_7). This mirrors the change from metallic to non-metallic character. Hydrides likewise change from ionic (with reactive metals) to covalent (with non-metals).

3.6 Position of Hydrogen

Hydrogen has one electron ($1s^1$). Like Group IA metals it can lose one electron to form H^+; like Group VIIA halogens it needs one electron to complete its shell and can form H^-. Because it resembles both families but is identical to neither, its position is described as anomalous, and it is usually placed separately at the top of the table.

4. Important Definitions

Term	Definition
Periodic table	A table in which elements are arranged in order of increasing atomic number in periods and groups.

Term	Definition
Modern periodic law	The properties of elements are a periodic function of their atomic numbers.
Periodicity	The regular repetition of properties of elements after certain regular intervals.
Period	A horizontal row in the periodic table; there are 7 periods.
Group	A vertical column in the periodic table; there are 18 groups.
Atomic radius	The distance from the nucleus to the outermost electron shell of an atom.
Ionization energy	The minimum energy required to remove the outermost electron from an isolated gaseous atom.
Electron affinity	The energy released when an electron is added to an isolated gaseous atom.
Electronegativity	The tendency of an atom to attract the shared electron pair in a covalent bond.
Shielding effect	The reduction in nuclear attraction on valence electrons caused by inner-shell electrons.
Effective nuclear charge	The net positive charge experienced by the valence electrons of an atom.

5. Formulas / Rules

Quick Rules to Remember

Effective nuclear charge: $Z_{eff} = Z - S$ (Z = atomic number, S = shielding constant)

Across a period (left → right): atomic & ionic radius decrease; IE, EA, EN increase.

Down a group (top → bottom): atomic & ionic radius increase; IE, EA, EN decrease.

Cation < parent atom < anion (in size).

Metallic character: decreases across a period, increases down a group.

Oxide nature across a period: basic → amphoteric → acidic.

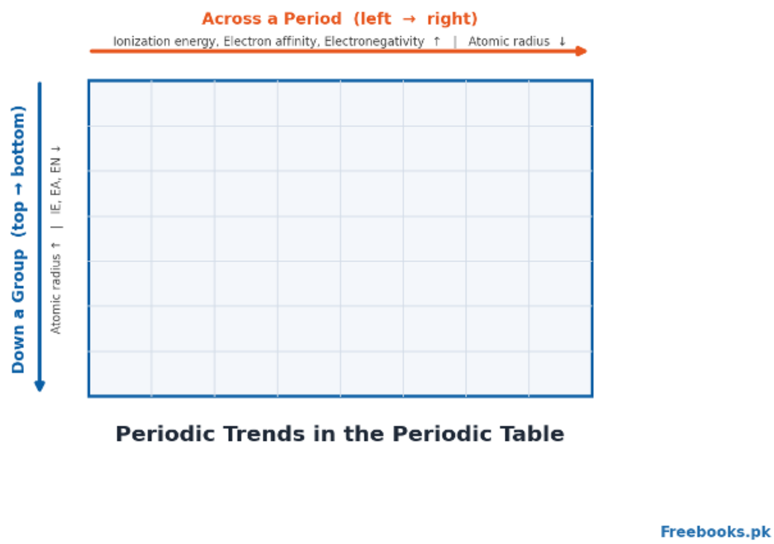

Figure 1.6 — Summary of periodic trends across a period and down a group.

6. Diagrams / Illustrations

The key diagrams for this chapter are collected here for quick revision. Study each one until you can redraw and label it from memory.

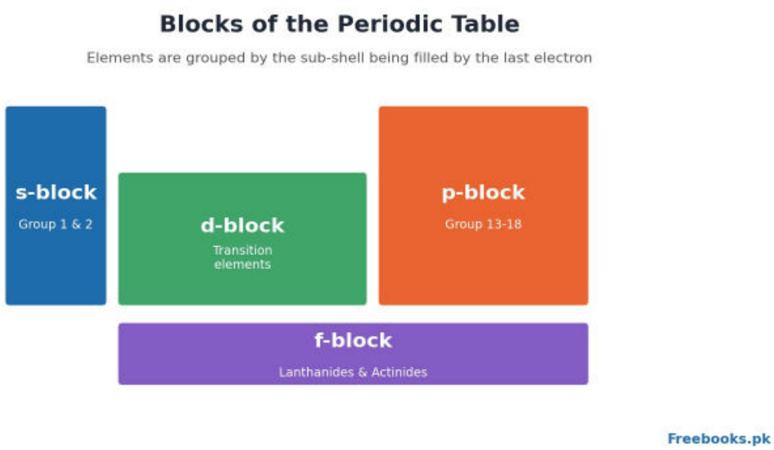

Diagram 1 — Blocks of the periodic table (s, p, d, f).

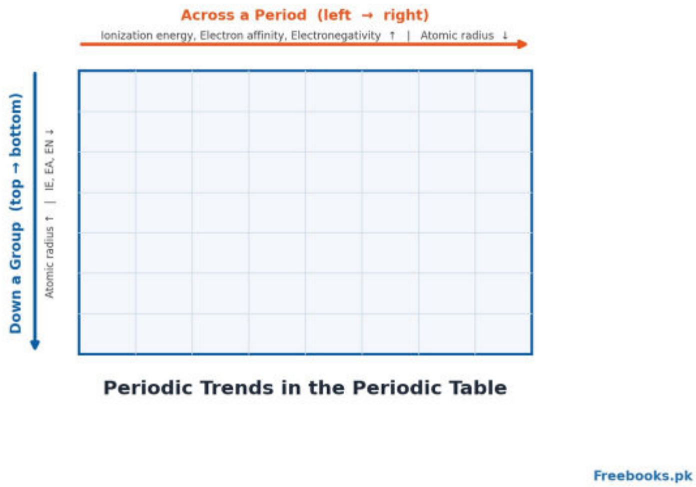

Diagram 2 — Direction of periodic trends.

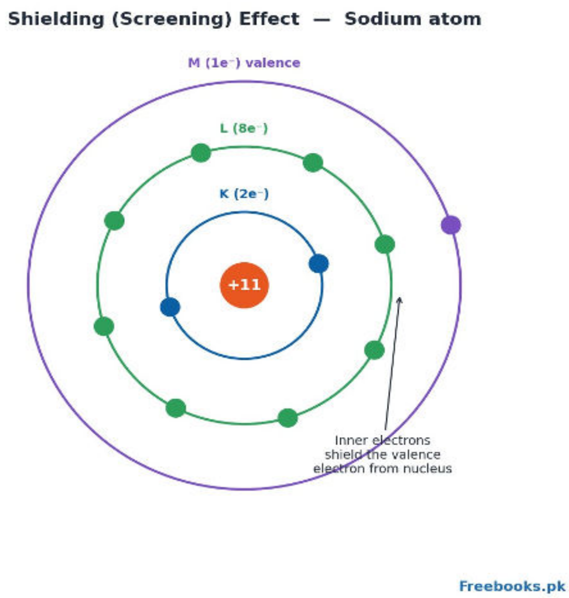

Diagram 3 — Shielding effect in a sodium atom.

7. Solved Examples

Example 1 — Comparing atomic size

Q: Which atom is larger, Na or Cl? Give a reason.

Solution: Both Na and Cl are in Period 3. Moving from Na to Cl the nuclear charge increases while electrons enter the same third shell, so the shell is pulled in more tightly. Therefore Na is larger than Cl.

Example 2 — Order of ionization energy

Q: Arrange Li, Na and K in increasing order of first ionization energy.

Solution: All three are Group IA metals. Down the group size increases and shielding increases, so the outer electron is removed more easily. Hence: K < Na < Li.

Example 3 — Nature of an oxide

Q: Predict whether the oxide of magnesium (MgO) is acidic or basic.

Solution: Magnesium is a metal (Group IIA). Metals form basic oxides, so MgO is basic — it reacts with acids to form a salt and water.

8. Short Questions

Q1. Define the modern periodic law.
Ans. The properties of elements are a periodic function of their atomic numbers.

Q2. What is a triad? Give one example.
Ans. A group of three elements with similar properties in which the middle element's atomic mass is the average of the other two, e.g. Li, Na, K.

Q3. Why does atomic radius decrease across a period?
Ans. Because nuclear charge increases while electrons enter the same shell, pulling it inward.

Q4. Why does ionization energy increase across a period?
Ans. Smaller size and higher effective nuclear charge hold the outer electron more tightly, so more energy is needed to remove it.

Q5. Why is a cation smaller than its parent atom?
Ans. It has lost electrons (often a whole shell), so the remaining electrons are held more tightly by the same nucleus.

Q6. Why do noble gases have almost zero electron affinity?
Ans. Their outermost shells are already complete, so they have no tendency to accept an extra electron.

Q7. Define shielding effect.
Ans. The decrease in nuclear attraction on valence electrons caused by the repulsion of inner-shell electrons.

Q8. Why is the position of hydrogen anomalous?
Ans. It resembles both Group IA (loses 1 e^-) and Group VIIA (gains 1 e^-), but is identical to neither.

9. Long Questions

Q1. Trace the development of the periodic table from Dobereiner to Moseley.

Dobereiner (1829) grouped elements into triads of three similar elements, where the middle element's atomic mass was the average of the other two. Newlands (1864) proposed the Law of Octaves, noticing that every eighth element repeated the properties of the first when arranged by increasing atomic mass, but this failed for heavier elements. Mendeleev (1869) arranged elements by increasing atomic mass in periods and groups, left gaps for undiscovered elements, and predicted their properties — a major success. However, his table had defects: an uncertain position for hydrogen, no place for isotopes, and a few anomalous pairs. Finally Moseley (1913) proved that atomic number, not atomic mass, is the true basis of classification, giving the modern periodic law and removing the anomalies.

Q2. Explain the variation of atomic radius, ionization energy and electronegativity across a period and down a group.

Across a period the nuclear charge (and effective nuclear charge) increases while electrons are added to the same shell. As a result the atomic radius decreases, and because the outer electrons are held more tightly, both ionization energy and electronegativity increase. Down a group a new shell is added at each step and shielding increases, so the atomic radius increases while ionization energy and electronegativity decrease because the outer electrons are farther from the nucleus and more shielded. These opposite directions are the reason fluorine (top-right region) is small and highly electronegative, whereas caesium (bottom-left) is large and gives up its electron easily.

Q3. What are effective nuclear charge and shielding effect? How do they control periodic properties?

Effective nuclear charge (Zeff) is the net positive charge felt by the valence electrons after the inner electrons partly cancel the pull of the nucleus (Zeff = Z − S). The shielding or screening effect is the repulsion of the outer electrons by the inner-shell electrons, which reduces the nuclear attraction they feel. When shielding is large, Zeff is small, the outer electrons are loosely held, the atom is larger, and ionization energy is low. When shielding is small (as across a period, where no new inner shell is added), Zeff is large, electrons are pulled in strongly, the atom is smaller, and ionization energy is high. Thus these two factors together explain nearly every trend in the chapter.

10. Multiple Choice Questions (MCQs)

1. The modern periodic law is based on:
(A) Atomic mass (B) Atomic number (C) Atomic radius (D) Density

2. Which scientist arranged elements in triads?
(A) Newlands (B) Mendeleev (C) Dobereiner (D) Moseley

3. The most electronegative element is:
(A) Oxygen (B) Chlorine (C) Fluorine (D) Nitrogen

4. Down a group, atomic radius:
(A) Increases (B) Decreases (C) Remains same (D) First increases then decreases

5. Which of the following is an amphoteric oxide?
(A) Na_2O (B) MgO (C) Al_2O_3 (D) SO_3

6. Elements in which block are called transition elements?
(A) s-block (B) p-block (C) d-block (D) f-block

7. Ionization energy across a period generally:
(A) Increases (B) Decreases (C) Stays constant (D) Becomes zero

8. A cation is always ___ than its parent atom:
(A) Larger (B) Smaller (C) Equal (D) Unpredictable

9. The number of periods in the modern periodic table is:
(A) 6 (B) 7 (C) 8 (D) 18

10. Noble gases have electron affinity that is nearly:
(A) Very high (B) Zero (C) Negative infinite (D) Same as halogens

MCQ Answer Key
1-B 2-C 3-C 4-A 5-C 6-C 7-A 8-B 9-B 10-B

11. Quick Revision / Summary

- Elements are arranged by increasing atomic number (Modern Periodic Law — Moseley).
- 7 periods (rows) and 18 groups (columns); four blocks: s, p, d, f.
- Periodicity = regular repetition of properties due to repeating valence-shell configuration.
- Across a period →: radius ↓, ionization energy ↑, electron affinity ↑, electronegativity ↑.
- Down a group ↓: radius ↑, ionization energy ↓, electron affinity ↓, electronegativity ↓.
- Cation < parent atom < anion (size).
- Oxides across a period: basic → amphoteric → acidic.
- Fluorine = most electronegative; hydrogen has an anomalous position.

12. Exam Tips

Score-Boosting Tips

Learn the two master ideas — effective nuclear charge and shielding — and you can reason out every trend instead of memorising them.

In trend questions, always mention BOTH the direction (across/down) AND the reason (size + Zeff/shielding).

Remember the small dips: B < Be and O < N in ionization energy — examiners love these exceptions.

Practise drawing and labelling the blocks diagram and the shielding diagram; labelled diagrams earn full marks.

For 'nature of oxide' questions, first decide metal vs non-metal, then say basic vs acidic (amphoteric for Al, Zn, Be).

Common confusion: ionization energy is energy absorbed; electron affinity is energy released — do not mix them.