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Classification of Elements and Periodicity in Properties Class 11 Notes | CBSE Chemistry Chapter 3

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Classification of Elements and Periodicity in Properties Class 11 Notes | CBSE Chemistry Chapter 3

Classification of Elements and Periodicity in Properties is Chapter 3 of CBSE Class 11 Chemistry — the chapter that turns a chaotic list of 118 elements into one beautifully organised map. Once you understand *why* the periodic table is built the way it is, you can predict the size, reactivity, and bonding behaviour of an element you have never even met. That single skill carries you through inorganic and physical chemistry for the next two years.

By the end of these notes you will be able to write the electronic configuration of any element, place it in the correct block, group, and period, and confidently explain how atomic radius, ionization enthalpy, electron gain enthalpy, and electronegativity change across a period and down a group. This is a high-weightage chapter carrying roughly 6–8 marks in boards, and the conceptual foundation for Chemical Bonding, the p-Block, d-Block, and coordination chemistry.

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Key Concepts

1. Genesis of Periodic Classification

Long before the modern table, chemists kept trying to group elements with similar properties so the subject would be easier to learn. Each attempt got a little closer to the truth.

- **Döbereiner's Triads (1829):** elements arranged in groups of three where the atomic mass of the middle element was roughly the average of the other two (e.g. Li, Na, K). Failed because very few such triads existed.
 - **Newlands' Law of Octaves (1865):** when elements were arranged by increasing atomic mass, every eighth element had properties similar to the first — like the eighth note of a musical octave. Worked only up to calcium.
 - **Lothar Meyer (1869):** plotted atomic volume against atomic mass and obtained a periodic curve, showing properties repeat at intervals.
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2. Mendeleev's Periodic Law

Dmitri Mendeleev (1869) gave the first widely accepted table. His **periodic law** states: *the physical and chemical properties of elements are a periodic function of their atomic masses.*

He arranged 63 known elements in order of increasing atomic mass into horizontal **periods** and vertical **groups** of similar elements.

Merits of Mendeleev's Table

- He left **gaps** for undiscovered elements and predicted their properties (eka-aluminium = gallium, eka-silicon = germanium) with remarkable accuracy.
- He placed elements by property even when it broke the mass order (e.g. he predicted corrected atomic masses).

Defects of Mendeleev's Table

- **Anomalous pairs:** some elements with higher atomic mass came before lower ones (Ar before K; Co before Ni).
 - Position of **isotopes** could not be explained (same element, different masses).
 - Hydrogen had no fixed position (resembles both Group 1 and Group 17).
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3. Modern Periodic Law and Present Form of the Table

Henry Moseley (1913) showed from X-ray studies that **atomic number (Z)**, not atomic mass, is the fundamental property of an element. This corrected all of Mendeleev's anomalies.

Modern Periodic Law: the physical and chemical properties of elements are a periodic function of their **atomic numbers**.

Features of the Long Form Table

- **7 periods (horizontal rows)** = number of principal energy shells; period number = value of n for the outermost shell.
 - **18 groups (vertical columns)** = elements with the same outer electronic configuration and similar properties.
 - The number of elements in each period (2, 8, 8, 18, 18, 32, 32) is fixed by how many electrons fill the available subshells ($2n^2$ rule).
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4. Electronic Configuration and Types of Elements (s, p, d, f Blocks)

The table is divided into four blocks based on the subshell into which the **last (differentiating) electron** enters.

Block	Groups	Outer Configuration	Examples
s-block	1 & 2	ns^{1-2}	Alkali & alkaline earth metals (Na, Ca)
p-block	13–18	$ns^2 np^{1-6}$	B, C, N, O, halogens, noble gases
d-block	3–12	$(n-1)d^{1-10} ns^{0-2}$	Transition metals (Fe, Cu, Zn)
f-block	placed below	$(n-2)f^{1-14}$	Lanthanoids & actinoids

- **s- and p-block** together form the **representative (normal) elements**.
- **d-block** = **transition elements**; **f-block** = **inner transition elements**.
- **Group 18** = noble gases ($ns^2 np^6$, fully filled, very stable).

Tip to locate an element: the period = highest value of n ; for s/p-block the group = (outer electrons) or (10 + outer electrons) for p-block; for d-block group = $(n-1)d + ns$ electrons.

5. Atomic and Ionic Radius

Atomic radius is the distance from the centre of the nucleus to the outermost shell of electrons. We use covalent, van der Waals, or metallic radii depending on the bonding.

- **Across a period (left → right):** atomic radius **decreases**, because nuclear charge increases while electrons enter the same shell, pulling them inward.
- **Down a group (top → bottom):** atomic radius **increases**, because a new shell is added in each period.

Ionic Radius

- A **cation** is smaller than its parent atom (it loses a shell / has higher effective nuclear charge per electron).
- An **anion** is larger than its parent atom (added electrons increase repulsion).
- **Isoelectronic species** (same number of electrons): radius decreases as nuclear charge increases. Order: $N^{3-} > O^{2-} > F^- > Na^+ > Mg^{2+} > Al^{3+}$.

6. Ionization Enthalpy (IE)

Ionization enthalpy is the minimum energy required to remove the most loosely bound electron from an isolated gaseous atom in its ground state.



- **Across a period:** IE generally **increases** (nuclear charge rises, size falls).
- **Down a group:** IE **decreases** (size increases, shielding increases).
- Successive ionization enthalpies always increase: $IE_1 < IE_2 < IE_3$.

Exceptions (very important for exams): Be has higher IE than B (stable fully filled $2s^2$), and N has higher IE than O (stable half-filled $2p^3$ configuration).

7. Electron Gain Enthalpy ($\Delta_e gH$)

Electron gain enthalpy is the energy change when an electron is added to an isolated gaseous atom to form an anion.



- More negative $\Delta_e gH$ means the atom releases more energy and accepts the electron more readily.
- **Across a period:** becomes more negative (atoms get smaller, more eager for electrons).
- **Down a group:** becomes less negative (size increases).
- **Halogens** have the most negative electron gain enthalpies.

Exception: $\Delta_e gH$ of **chlorine is more negative than fluorine**, because F is so small that incoming electrons face high repulsion in its compact $2p$ shell.

8. Electronegativity

Electronegativity is the tendency of an atom in a molecule to attract the shared pair of electrons towards itself. Unlike IE and $\Delta_e gH$, it is a relative property with no units.

- **Across a period:** electronegativity **increases** (size decreases, nuclear pull increases).
- **Down a group:** electronegativity **decreases**.
- **Fluorine** is the most electronegative element (Pauling value 4.0).

Electronegativity decides bond polarity, the metallic/non-metallic character, and the acidic/basic nature of oxides.

9. Valence (Valency)

Valence is the combining capacity of an element, decided by the number of valence electrons.

- **Across a period:** valence towards hydrogen rises $1 \rightarrow 4$ then falls $4 \rightarrow 1$ (CH_4 , NH_3 , H_2O , HF).
 - **Down a group:** valence usually stays constant (same outer configuration).
 - For oxides, the highest valence often equals the group number for representative elements.
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10. Periodic Trends in Chemical Properties

The physical trends above translate directly into chemical behaviour.

- **Metallic character** increases down a group and decreases across a period (metals lose electrons easily — low IE).
- **Non-metallic character** increases across a period and decreases down a group.

- **Nature of oxides:** metallic oxides are basic (Na_2O), non-metallic oxides are acidic (SO_3 , Cl_2O_7), and oxides of borderline elements are amphoteric (Al_2O_3 , ZnO).
- Reactivity of **metals increases** down a group; reactivity of **non-metals decreases** down a group.

[DIAGRAM: A periodic table with arrows — atomic radius and metallic character increasing down/left; ionization enthalpy, electronegativity and non-metallic character increasing up/right towards fluorine.]

11. Diagonal Relationship

Some elements of period 2 resemble the period-3 element placed diagonally to the right (Li–Mg, Be–Al, B–Si). This happens because the diagonal pair has similar charge/size ratios and similar electronegativities.

12. Nomenclature of Elements with $Z > 100$

To avoid naming disputes for newly synthesised superheavy elements, IUPAC gives a temporary systematic name based on the digits of the atomic number, using numerical roots, with the suffix **-ium**.

Digit	0	1	2	3	4	5	6	7	8	9
Root	nil	un	bi	tri	quad	pent	hex	sept	oct	enn

- $Z = 101 \rightarrow$ **Unnilunium (Unu), Md**
- $Z = 104 \rightarrow$ **Unnilquadium (Unq), Rf**
- $Z = 120 \rightarrow$ **Unbinilium (Ubn)**

Rule: write the roots for each digit in order, add **-ium**, and the symbol is the first letter of each root. If "i" of "bi"/"tri" is followed by "ium", one "i" is dropped; "enn" + "nil" drops one "n".

Weightage in Board & Entrance Exams

Exam	Typical Weightage	Most-Tested Areas
CBSE Board (Class 11)	6–8 marks	Periodic trends, IE exceptions, electronic configuration, blocks
JEE Main / Advanced	1–2 questions	Isoelectronic radii, IE order, electronegativity, nomenclature
NEET	2–3 questions	Atomic/ionic radius, IE & Δ_{egH} exceptions, metallic character

[TABLE: Question-type split — VSA (1 mark): definitions, blocks, most electronegative element; SA (2–3 marks): trend reasoning, IE exceptions; LA (5 marks): compare across period/group, predict properties, nomenclature.]

Important Definitions

Term	Definition
Modern periodic law	Properties of elements are a periodic function of their atomic numbers
Period	Horizontal row; equals the number of the outermost principal shell (n)
Group	Vertical column of elements with the same outer electronic configuration
Atomic radius	Distance from the nucleus to the outermost electron shell
Isoelectronic species	Atoms/ions with the same number of electrons (e.g. O^{2-} , F^- , Na^+)
Ionization enthalpy	Energy to remove the outermost electron from a gaseous atom
Electron gain enthalpy	Energy change when an electron is added to a gaseous atom
Electronegativity	Tendency of an atom to attract a shared electron pair in a bond
Representative elements	s- and p-block elements (Groups 1, 2 and 13–18)
Diagonal relationship	Similarity between an element and the one diagonally right below it

Solved Examples

Example 1

An element has the electronic configuration $1s^2 2s^2 2p^6 3s^2 3p^3$. Identify its block, group, and period.

Answer: Last electron enters the 3p subshell $\rightarrow$ **p-block**. Highest $n = 3 \rightarrow$ **Period 3**. Outer electrons = $2 + 3 = 5 \rightarrow$ group = $10 + 5 =$ **Group 15**. The element is phosphorus ($Z = 15$).

Example 2

Arrange the isoelectronic species N^{3-} , O^{2-} , F^- , Na^+ , Mg^{2+} in increasing order of ionic radius.

Answer: All have 10 electrons; radius decreases as nuclear charge rises. Increasing order: **$Mg^{2+} < Na^+ < F^- < O^{2-} < N^{3-}$** .

Example 3

Why is the first ionization enthalpy of nitrogen greater than that of oxygen?

Answer: Nitrogen has a stable, exactly half-filled $2p^3$ configuration which is hard to disturb, so removing an electron needs more energy. Oxygen ($2p^4$) loses one electron to reach the stable half-filled state, so its IE_1 is lower.

Example 4

The electron gain enthalpy of chlorine is more negative than that of fluorine. Explain.

Answer: Fluorine is very small, so its compact 2p shell causes strong electron–electron repulsion when an extra electron is added, reducing the energy released. Chlorine is larger (3p shell), repulsion is less, so its Δ_{egH} is more negative.

Example 5

Which of Be and B has the higher first ionization enthalpy, and why?

Answer: Beryllium. Be has a stable fully filled $2s^2$ configuration, so its outer electron is tightly held. Boron's single 2p electron is at higher energy and easier to remove, so B has a lower IE_1 .

Example 6

Write the IUPAC systematic name and symbol for the element with atomic number 105.

Answer: Digits 1-0-5 $\rightarrow$ un-nil-pent + ium = **Unnilpentium**, symbol **Unp** (now named Dubnium, Db).

Important Questions for Board Exams

1-Mark Questions (VSA)

1. State the modern periodic law.
2. Name the most electronegative element in the periodic table.
3. Which block do transition elements belong to?
4. Why is the size of a cation smaller than its parent atom?
5. Write the IUPAC name of the element with $Z = 120$.

2–3-Mark Questions (SA)

1. Why does atomic radius decrease across a period but increase down a group? Explain.
2. Explain why the first ionization enthalpy of N is greater than that of O.
3. Arrange O^{2-} , F^- , Na^+ , Mg^{2+} in increasing order of size and justify.
4. Distinguish between ionization enthalpy and electron gain enthalpy with one example each.

5-Mark Questions (LA)

1. Discuss the variation of atomic radius, ionization enthalpy, electron gain enthalpy, and electronegativity across a period and down a group, with reasons.
2. What are the defects of Mendeleev's periodic table, and how does the modern periodic law remove them?

3. Explain the trend in metallic and non-metallic character and the acidic/basic nature of oxides across period 3.

Quick Revision Points

- Döbereiner triads → Newlands octaves → Mendeleev (atomic mass) → Modern law (atomic number)
 - Modern periodic law: properties are a periodic function of atomic number Z
 - Long form: 7 periods, 18 groups; period = outermost shell n
 - Blocks: s (Gp 1–2), p (Gp 13–18), d (Gp 3–12, transition), f (inner transition)
 - Atomic radius: decreases across a period, increases down a group
 - Cation < atom < anion; isoelectronic radius falls as nuclear charge rises
 - IE: increases across a period, decreases down a group; exceptions Be > B, N > O
 - Δ_{egH} : most negative for halogens; exception Cl more negative than F
 - Electronegativity: increases across a period, decreases down a group; F is highest (4.0)
 - Metallic character ↓ across, ↑ down; oxides: metal = basic, non-metal = acidic, borderline = amphoteric
 - $Z > 100$ nomenclature: nil-un-bi-tri-quad-pent-hex-sept-oct-enn + ium
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Next Chapter: Chapter 4 — Chemical Bonding and Molecular Structure



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