4. Periodic Classification of Elements

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🎯 Learning outcomes

After the completion of the chapter, the students should be able to:

- ✓ Explain the modern periodic table and its features.
- ✓ Classify the elements of the periodic table in different blocks and groups.
- ✓ Identify the elements as metals, non-metals, and metalloids.
- ✓ Define the term nuclear charge and effective nuclear charge.
- ✓ Explain and interpret the Periodic trend of atomic radii, ionic radii, ionisation energy, electronegativity, electron affinity, and metallic characters of elements.

🧠 Initial assessment- What do you already know?

1. What was the main basis for the classification of elements in Mendeleev's periodic table?
2. State Mendeleev's periodic law.
3. List three advantages and three limitations of Mendeleev's periodic table.

Introduction

The study of elements revealed that their physical and chemical properties show a regular pattern when arranged systematically. This led to the development of the **periodic table**, one of the most important tools in chemistry. It allows scientists to understand relationships between elements and predict their properties.

Modern Periodic Law and Modern Periodic Table

Early attempts at classification, such as Mendeleev's periodic table, were based on **atomic mass**. However, some inconsistencies remained unresolved. In 1913, **Henry Moseley** established that the **atomic number (Z)**, rather than atomic mass, is the fundamental property governing periodicity.

> **Key Term: Modern Periodic Law**
>
> **The physical and chemical properties of elements are periodic functions of their atomic numbers.**
>
> When elements are arranged in increasing order of atomic number, their properties repeat at regular intervals.

The modern periodic law overcomes the defects of Mendeleev's periodic table due to the following reasons:

1. Since the atomic number is the basis for classification, isotopes do not need separate positions.
2. Position of Ar before K, position of Ni before Co, position of Te before I are justified since the former has a lower atomic number than the latter in each case.

Main features of the modern periodic table:

The modern periodic table is based on the **electronic configuration** of elements. Its main structural features are as follows:

1. There are seven rows, which are called periods and indicated by 1, 2, 3, 4, 5, 6, and 7.
2. In a period, the outermost electronic configuration changes gradually, which causes the gradual change in properties. In the next period, a similar outer electronic configuration reappears. Hence, the properties also repeat.
3. There are 18 columns, which are called groups. Each group consists of a set of elements having the same outer electronic configuration; hence, they show similar properties. For example, Li, Na, and K have the same outer electronic configuration of ns^1, and they are placed under group 1.

4. Dissimilar elements do not fall in the same group since subgroups have been separated.
5. The modern periodic table completely separates metals from non-metals.
6. The left-hand side of the table consists of highly reactive metals, and the right-hand side consists of non-metals; the middle portion contains transition elements.
7. Elements have been divided into four distinct blocks: s-block, p-block, d-block, and f-block.
8. Lanthanides and actinides are not placed in the main body of the periodic table.

Advantages of the Modern Periodic Table

1. **Explains periodicity:** Repeating electronic configurations explain why properties recur at regular intervals.
2. **Resolves Mendeleev's anomalies:** Atomic number ordering justifies the placement of Ar before K, Co before Ni, and Te before I pairs.
3. **No problem with isotopes:** Isotopes have the same Z, so they automatically share one position.
4. **Clear separation of metals and non-metals:** The diagonal boundary(from boron to astatine) makes classification straightforward.
5. **Block division:** s, p, d, f blocks allow systematic study of chemical properties.

Limitations of the Modern Periodic Table

- **Position of hydrogen:** Hydrogen shares properties with both alkali metals (Group 1) and halogens (Group 17). It does not fit perfectly into either group.
- **Position of helium:** Helium's outermost configuration ($1s^2$) resembles Group 2, but it is placed in Group 18 due to its chemical inertness. Also, it lacks a p-orbital, unlike its p-block neighbours.
- **Lanthanides and actinides:** Placing the f-block elements at the bottom of the table (outside the main body) technically violates the continuous arrangement expected by the periodic law.

Conclusion

The modern periodic table provides a systematic framework for understanding the properties of elements. By studying periodic trends and electronic configurations, chemists can predict the behavior, reactivity, and bonding patterns of elements with great accuracy.

Classification of elements into different groups, periods, and blocks

- There are seven rows, which are called periods and indicated by 1, 2, 3, 4, 5, 6, and 7.
- In a period, the outermost electronic configurations change gradually, which causes the gradual change in properties. In the next period, a similar outer electronic configuration reappears. Hence, the properties also repeat.
- There are 18 columns, which are called groups. Each group consists of a set of elements having the same outer electronic configuration; hence, they show similar properties. For example, Li, Na, and K have the same outer electronic configuration of $ns^{1,}$ and they are placed under group 1.
- Dissimilar elements do not fall in the same group since subgroups have been separated.
- Elements have been divided into four distinct blocks: s-block, p-block, d-block, and f-block.

Properties of s-Block elements (contain s-electrons in valence shell)

Valence configuration: ns^{1-2}

1. All are active metals except hydrogen.
2. Oxidation states: +1 (Group 1) and +2 (Group 2).
3. They form basic oxides and hydroxides.
4. They impart characteristic colour to the flame.
5. Generally, they form ionic salts with nonmetals.
6. low ionization potentials.
7. very small electron affinities.
8. Solids at room temperature.
9. They are good reducing agents.

Properties of p-Block elements:

1. Valence electronic configuration: $ns^2\ np^{1-6}$
2. Mostly non-metals; also contains metalloids and some metals.
3. They have variable oxidation states.
4. They form acidic oxides(except group 18).
5. Generally, they form covalent compounds. Halogens form salts with alkali metals
6. They have high ionization potentials.
7. They have large electron affinities.
8. They may be solids, liquids, or gases at room temperature (Br is a liquid).

Properties of d-Block elements: (contain d-electrons in valence shell).

1. general electronic configuration : $(n-1)d^{1-10}ns^{0-2}$.
2. Lie between s- and p-block, showing "transitional" properties.
3. Almost all are metals (dense, hard, high-melting).
4. mostly variable oxidation states due to partly filled d-orbitals. (Except Sc, Zn, Cd, etc.)
5. Many compounds are brightly coloured — a key identification feature.
6. Many exhibit paramagnetic, ferromagnetic behavior.
7. Excellent catalysts (e.g., Fe in the Haber process, Ni in hydrogenation).

Properties of f-Block elements: (contain f-electrons in valence shell)

1. Electronic configuration: ns^2 (n-) d^{0-1}(n-2) f^{1-14}

2. Also called inner transition elements. They are: Lanthanides and actinides.

Lanthanides ($_{58}$Ce- $_{71}$Lu)

3. All lanthanides closely resemble lanthanum in properties and are chemically similar to each other.

5. Mostly trivalent. (Except for cerium (III and IV) and europium (III and II))

6. Most lanthanides are widely used in lasers and magnets

7. Commonly used in sunglasses — they absorb UV and IR radiation

Actinides ($_{90}$Th– $_{103}$Lr)

8. All actinides are radioactive.

9. These are highly electropositive (show +3, to +6 oxidation states)

10. These metals tarnish in the air.
11. They have several isotopes.
12. They react with boiling water or dilute acids to give H_2 gas.
13. These directly combine with non-metals.

IUPAC classification of elements

IUPAC (International Union of Pure and Applied Chemistry) numbers the main-body groups from 1 to 18. This replaces the older system that used Roman numerals with A/B suffixes. The table below maps the two systems:

IUPAC	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
old	IA	IIA	IIIB	IVB	VB	VIB	VIIB	VIII			IB	IIB	IIIA	IVA	VA	VIA	VIIA	0

Nuclear charge and effective nuclear charge

Nuclear Charge (Z)

The total positive charge on the nucleus of an atom, equal to the number of protons. For example, the nuclear charge of sodium (Na, Z = 11) is +11.

In a multi-electron atom, the inner electrons partially block (or "shield") the outer electrons from feeling the full attraction of the nucleus. Because of this, the outermost electrons experience a reduced, effective attraction.

Key Term: Effective Nuclear Charge (Zeff)
The net positive charge actually experienced by a valence electron after accounting for the shielding effect of inner electrons. $\mathbf{Z_{eff} = Z - \sigma}$, where σ is the shielding (screening) constant.

Periodic Trend and Periodicity

Periodic Properties of Elements

The properties below show a regular, predictable variation across periods and down groups. Understanding these trends allows us to predict and compare element behaviour without memorizing each case separately.

1 Atomic radius

An atom is considered to be spherical. Hence, the size of an atom is expressed in terms of atomic radius.

Key Term: Atomic Radius

The distance between the centre of the nucleus and the outermost electron shell of an atom.

It can be expressed as covalent radius (half the distance between two bonded atoms of the same element), metallic radius, or Van der Waals radius.

- Variation of atomic radius across a period

The nuclear charge increases, and an extra electron is added to the same shell. Hence, atomic size decreases with increasing atomic number in a period.

- Variation of atomic radius down a group

The nuclear charge increases, which means the size should decrease, but an electron is added to the new shell away from the nucleus, resulting in an increase in the size of the atom. The latter effect outweighs the former. Hence, atomic size increases down the group.

On going down in a group, the atomic radius increases because of the increase in an extra shell.

Ionic radius

Ionic radius may be defined as the distance between the centre of the nucleus and a point at which the nucleus has its influence on its electron in an ion.

1. Cations are smaller than the parent atom.

This is generally due to the removal of the valence shell and an increase in the effective nuclear charge.

2. Anions are larger than the parent atom.

This can be explained on the basis of a decrease in the effective nuclear charge in an anion.

3. Isoelectronic Species

Ions or atoms having the same number of electrons but different nuclear charges are called isoelectronic species. Their size decreases with an increase in the nuclear charges.

2 Ionization Energy/Ionization Potential (IE/IP)

Key Term: Ionisation Energy

The minimum amount of energy required to remove the most loosely held electron from an isolated gaseous atom in its ground state. The atom becomes a cation.

$$M(g) + IE \longrightarrow M^+(g) + e^-$$

- Successive Ionization energies

It is possible to remove more than one electron from an atom successively.

The minimum amount of energy required to remove the most loosely held electron from an isolated gaseous atom in its ground state is called the first ionization energy of that element.

$$M(g) + IE_1 \longrightarrow M^+(g) + e^-$$

The minimum amount of energy required to remove the most loosely held electron from an isolated gaseous monopositive ion in its ground state is called the second ionization energy of that element.

$$M(g) + IE_2 \longrightarrow M^+(g) + e^-$$

The minimum amount of energy required to remove the most loosely held electron from an isolated gaseous dipositive ion in its ground state is called the third ionization energy of that element.

$M(g) + IE_3 \longrightarrow M^+(g) + e^-$

For an element, $IE_1 < IE_3 < IE_3$

- Factors affecting Ionization energy

The ionization energy of an electron depends upon the following factors.

a) **Atomic size**: As the size of the atom decreases, electrons are more strongly attracted by the nucleus, and more energy is required to remove the electron; hence, the greater the IE.

b) **Nuclear charge**: As the nuclear charge increases, electrons are more strongly attracted by the nucleus, and more energy is required to remove the electron. Hence, the greater the IP.

c) **Electronic configuration**: Half-filled and fully filled orbitals are more stable than other configurations; hence, the IE is higher in such cases. For example, Beryllium has a higher IE than Boron for this reason.

d) **Screening or Shielding effect:** In multi-electron atoms, in addition to attraction between the nucleus and the electrons, there is repulsion between the electrons, so that the valence electron experiences less nuclear attraction. This effect of inner electrons is called the screening effect or shielding effect. The greater the number of electrons, the higher the shielding effect. As the screening effect increases, electrons are less strongly attracted by the nucleus, and less energy is required to remove an electron. Hence, ionization energy decreases with an increase in screening effect.

- Variation of IE within a group

On moving from top to bottom in a group, although nuclear charge increases, the screening effect increases, and the size of atoms increases, due to which electrons are less strongly attracted and less energy is required to remove an electron. Hence, ionization energy decreases.

- Variation of IP over a period

On moving from left to right in a period, nuclear charge increases, and the size of the atom decreases, due to which electrons are more strongly attracted by the nucleus and more energy is required to remove the electron. Hence, ionization energy increases.

Property	Period (Left → Right)	Group (Top → Bottom)
Ionisation Energy	Increases → (Z↑, size↓, stronger pull)	Decreases ↓ (size↑, shielding↑, weaker pull)

3 Electron Affinity (EA)

Key Term: Electron Affinity

The amount of energy released when an electron is added to an isolated gaseous atom in its ground state.

A high EA means the atom strongly attracts an additional electron.

- Main Factors affecting EA

a) Atomic size

The smaller the size of an atom, the greater the attraction between the electron in the outermost shell and the nucleus, hence the greater the electron affinity.

b) Nuclear charge

The more the nuclear charge in the atom, the greater the attraction between the electron in the outermost shell and the nucleus, thus increasing the EA.

c) Electronic configurations

Half-filled and fully filled orbitals are more stable than other configurations; hence, the EA is less in such cases.

- Variation of electron affinity in a period

On moving from left to right in a period, nuclear charge increases, and the size of the atom decreases, due to which the tendency of the atom to attract incoming electrons increases. Hence, electron affinity increases. However, there are irregularities in the general trend, mainly due to electronic configurations.

- Variation of electron affinity in a period

On moving from top to bottom in a group, although the nuclear charge increases, the screening effect increases as the size of the atom increases, due to which the tendency of the atom to attract incoming electrons decreases.

4 Electronegativity

Key Term: Electronegativity

The tendency of an atom in a molecule to attract the shared pair of electrons towards itself.

Unlike ionisation energy and electron affinity, it is NOT a property of an isolated atom — it only makes sense in the context of a chemical bond.

Variation of electronegativity in a period:

On moving from left to right in a period, nuclear charge increases, and the size of the atom decreases, due to which the tendency of the atom to attract a shared pair of electrons increases. Hence, electronegativity increases. e.g.

Li	Be	B	C	N	O	F	Ne
1.0	1.5	2.0	2.5	3.0	3.5	4.0	0

Variation of electronegativity in a group:

On moving from top to bottom in a group, although nuclear charge increases, screening effect increases, and size of the atom increases, due to this, electronegativity decreases down the group. e.g.

Fluorine	4.0
Chlorine	3.0
Bromine	2.8
Iodine	2.5

Property	Period (Left → Right)	Group (Top → Bottom)
Electronegativity	Increases → (Z↑, size↓, greater pull on shared e^-)	Decreases ↓ (size↑, shielding↑)

5 Metallic Characters (to be studied after metallic bonding)

The tendency of an element to lose electron(s) to form a cation is the characteristic of a metal. The elements having low ionization energy and low electronegativity (electropositive) show such characteristics. The elements on the left and lower left side of the periodic table are such elements.

Variation of metallic character in a period:
On moving from left to right of the periodic table, the ionization energies of elements decrease, hence the metallic character also decreases gradually.

Na > Mg > Al

Variation of metallic character in a group:
As moving down in a group of the periodic table, ionization energy decreases, hence metallic character increases.

Cs > Rb > K > Na > Li

Property	Period (Left → Right)	Group (Top → Bottom)
Metallic Character	Decreases → (IE increases, harder to lose e^-)	Increases ↓ (IE decreases, easier to lose e^-)

Did You Know?
Caesium (Cs) is the most electropositive (metallic) element. It is so reactive that it ignites spontaneously in air and explodes on contact with water!

Quick Summary of Periodic Trends:

Property	Across Period (→)	Down Group (↓)	Highest
Atomic Radius	Decreases	Increases	Cs
Ionisation Energy	Increases	Decreases	He
Electron Affinity	Increases*	Decreases	Cl
Electronegativity	Increases	Decreases	F
Metallic Character	Decreases	Increases	Cs

** Electron affinity shows exceptions at Groups 2, 5, and 18 due to stable electronic configurations.*

Key Points to Remember

1. The modern periodic law states that properties are periodic functions of atomic number (Z).
2. The periodic table has 7 periods and 18 groups. Elements in the same group have the same valence electron configuration and similar properties.
3. Elements are divided into s, p, d, and f blocks based on which sub-shell the last electron enters.
4. Effective nuclear charge (Zeff) = Z − σ. It increases across a period and drives many periodic trends.
5. Atomic radius decreases across a period; increases down a group.
6. Ionisation energy, electron affinity, and electronegativity generally increase across a period and decrease down a group.
7. Metallic character decreases across a period and increases down a group.