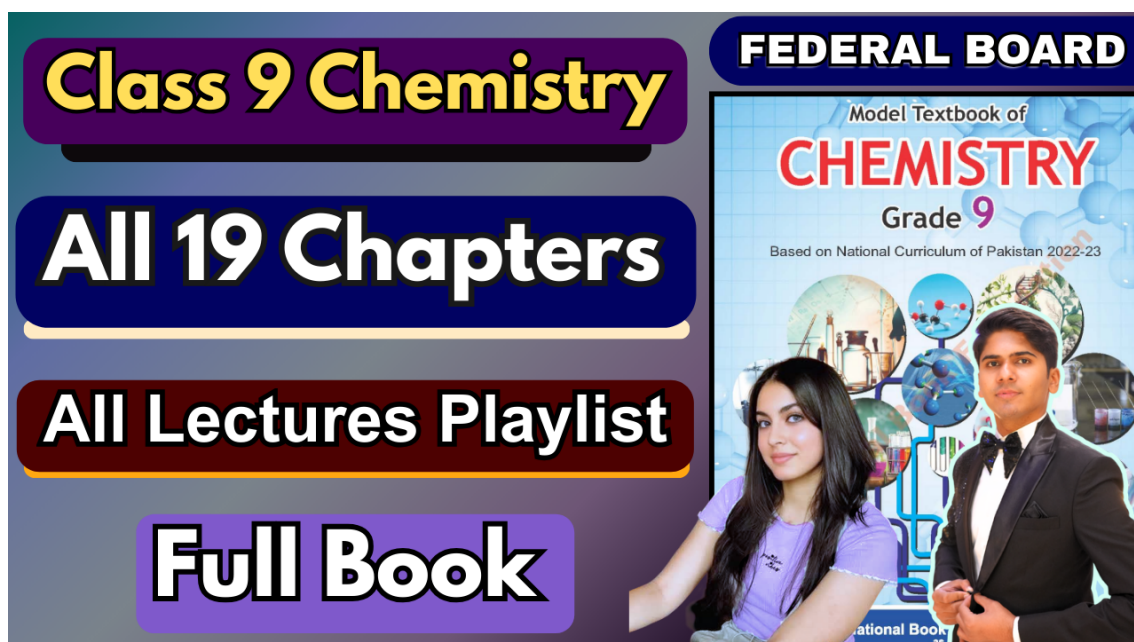


# Chapter 4: Periodic Table and Periodicity of Properties

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Only 23 elements were known until the end of the 18<sup>th</sup> century, to its development of 118 elements today. It is very difficult and impossible to remember information about the reactions, properties, and atomic masses of elements. So, we obviously need a way to organize our information about them. The periodic table is one of the most important tools in chemistry. It is very useful for understanding and predicting the properties of elements. For example, if you know the physical and chemical properties of one element in a group, you can predict the physical and chemical properties of any other element in the same group. The periodic table allows you to relate the reactivity tendencies of elements to their atomic structure. You can also predict which elements can form ionic or covalent bonds.

## 4.1. Periodic Table

After the discovery of atomic number by Moseley in 1913, it was noticed that atomic number could serve as a base for systematic arrangement of elements. Thus, elements are arranged in the order of increasing atomic number. A table showing systematic arrangement of elements is called a periodic table. It is based on the **Periodic law** that states if the elements are arranged in the order of their increasing atomic numbers, their properties are repeated in a periodic manner.

### 4.1.1. Periods and Groups of Elements

- **Periods** are the **horizontal rows** of the periodic table, and there are **7** of them. Elements in a period have a **different** number of valence electrons but the **same** number of shells. Properties change gradually across a period.
- **Groups** are the **vertical columns** of the periodic table, and there are **18** of them. Elements in a group have the **same** number of valence electrons (mostly) and **similar** chemical properties.

## 4.2. Periodic Trends

### 4.2.1. Atomic Size/Radius

Atomic size is the distance from the center of the nucleus to the outermost shell of an atom.

- **Across a Period (Left → Right):** Atomic size **decreases** due to the increase in nuclear charge, which pulls the valence electrons closer to the nucleus.
- **Down a Group (Top → Bottom):** Atomic size **increases** due to the addition of a new shell, which causes the **screening/shielding effect**.

### 4.2.2. Ionic Radius

The radius of an ion (charged atom).

- **Cation ( $X^+$ )** is **smaller** than its neutral atom because of the loss of the valence shell or greater nuclear charge acting on the remaining electrons.
- **Anion ( $X^-$ )** is **larger** than its neutral atom because of the increase in the number of electrons, which increases the repulsion and expands the size.
- **Isoelectronic Series:** A series of ions/atoms having the same number of electrons. The size decreases with increasing nuclear charge.

### 4.2.3. Ionization Energy (IE)

The minimum amount of energy required to remove the most loosely held electron from the outermost shell of a gaseous atom.

- **Across a Period (Left → Right):** Ionization energy **increases** due to the decrease in atomic size and increase in nuclear charge.
- **Down a Group (Top → Bottom):** Ionization energy **decreases** due to the increase in atomic size and the **shielding effect**.

### 4.2.4. Electron Affinity (EA)

The energy released when an electron is added to the valence shell of a gaseous atom.

- **Across a Period (Left → Right):** Electron affinity becomes **more negative** (more energy released) because the smaller size and greater nuclear charge mean the atom has a stronger attraction for the incoming electron.
- **Down a Group (Top → Bottom):** Electron affinity becomes **less negative** because the increased size and shielding effect mean the atom has a weaker attraction for the incoming electron.

### 4.2.5. Electronegativity (EN)

The ability of an atom to attract the shared pair of electrons towards itself in a covalent bond.

- **Across a Period (Left → Right):** Electronegativity **increases** due to the increase in nuclear charge and decrease in atomic size.

- **Down a Group (Top → Bottom):** Electronegativity **decreases** due to the increase in atomic size and the shielding effect.
- **Fluorine (F)** is the most electronegative element (Electronegativity = 4.0).

### 4.3. Electronegativity and Chemical Bonds

The difference in electronegativity between two atoms determines the type of chemical bond formed between them.

- **Non-Polar Covalent Bond:** The difference in electronegativity is **zero or very small** ( $\leq 0.4$ ). The electrons are shared equally.
- **Polar Covalent Bond:** The difference in electronegativity is **intermediate** ( $0.4 < \text{EN difference} < 1.7$ ). The electrons are shared unequally.
- **Ionic Bond:** The difference in electronegativity is **large** ( $\geq 1.7$ ). The electrons are completely transferred from one atom to another.

### 4.4. Valence Shell Electron Pair Repulsion (VSEPR) Theory

The VSEPR theory is used to predict the **geometry/shape** of simple molecules. **Principle:** Electron pairs (both bonding and lone pairs) around a central atom repel each other. They arrange themselves in a space to be as far apart as possible to **minimize the repulsion**.

- **Repulsion Order:**

1. Lone pair - Lone pair repulsion (strongest)
2. Lone pair - Bonding pair repulsion (intermediate)
3. Bonding pair - Bonding pair repulsion (weakest)

#### Examples of Molecular Shapes

- **Beryllium Chloride ( $\text{BeCl}_2$ ):** 2 Bonded Pairs, 0 Lone Pairs → **Linear** ( $180^\circ$ )
- **Boron Trifluoride ( $\text{BF}_3$ ):** 3 Bonded Pairs, 0 Lone Pairs → **Trigonal Planar** ( $120^\circ$ )
- **Methane ( $\text{CH}_4$ ):** 4 Bonded Pairs, 0 Lone Pairs → **Tetrahedral** ( $109.5^\circ$ )
- **Ammonia ( $\text{NH}_3$ ):** 3 Bonded Pairs, 1 Lone Pair → **Trigonal Pyramidal** ( $107.5^\circ$ )
- **Water ( $\text{H}_2\text{O}$ ):** 2 Bonded Pairs, 2 Lone Pairs → **Bent/V-shaped** ( $104.5^\circ$ )

### 4.5. Chemical Properties of Elements

#### 4.5.1. Metals

- Tendency to **lose electrons** easily.
- Form **positive ions** (cations).
- Reactivity **increases down a group** and **decreases across a period**.

- Example: Alkali metals (Group 1) are highly reactive.

#### 4.5.2. Non-Metals

- Tendency to **gain electrons** easily.
- Form **negative ions** (anions).
- Reactivity **decreases down a group** and **increases across a period**.
- Example: Halogens (Group 17) are highly reactive.

#### 4.6. Transition Elements

- Located in the **d-block** (Groups 3-12).
- Form coloured compounds, have variable valencies (oxidation states), and often act as catalysts.
- Examples: Iron (Fe), Copper (Cu), Gold (Au).

#### 4.7. Metalloids

- Elements with properties **between metals and non-metals**.
- Act as **semiconductors** (important in electronics).
- Found along the staircase line (e.g., Boron (B), Silicon (Si), Germanium (Ge)).

#### 4.8. Noble Gases (Group 18)


- Chemically **inert/unreactive**.
- Have a complete outermost shell (He: 2, others: 8).
- Used in lighting, welding, and industrial processes.
- Their non-reactive nature makes them suitable for **safe, stable environments** such as gas fillings and protective atmospheres.

#### 4.9. Comparison Of General Physical Properties Of Metals And Non-Metals

- **Thermal Conductivity:** Metals generally have high thermal conductivity, which means they can conduct heat easily. On the other hand, non-metals tend to have poor conductivity, making them less efficient at conducting heat.
- **Electrical Conductivity:** Metals are good conductors of electricity, because they have free electrons that can move freely in the metal lattice. Non-metals, with few exceptions such as graphite, are poor conductors of electricity because they lack free electrons.
- **Adaptability:** The metals are malleable and ductile. So, they can be hammered, drawn into wires or transformed into thin sheets without breaking. This property is due to metallic bonds which allow atoms exchange easily under pressure. Non-metals are neither malleable nor ductile, rather they are brittle.

- **Melting Points and Boiling Points:** Metals generally have high melting points and boiling points due to strong metallic bonds that require a lot of energy to break. On the other hand non-metals often have lower melting points and boiling points because their atoms and molecules are held by weaker bonds such as covalent bonds, van der Waals bonds, or hydrogen bonds that require less energy to break.





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