# Chapter 2: Matter

**All Lectures Uploaded on YouTube:**

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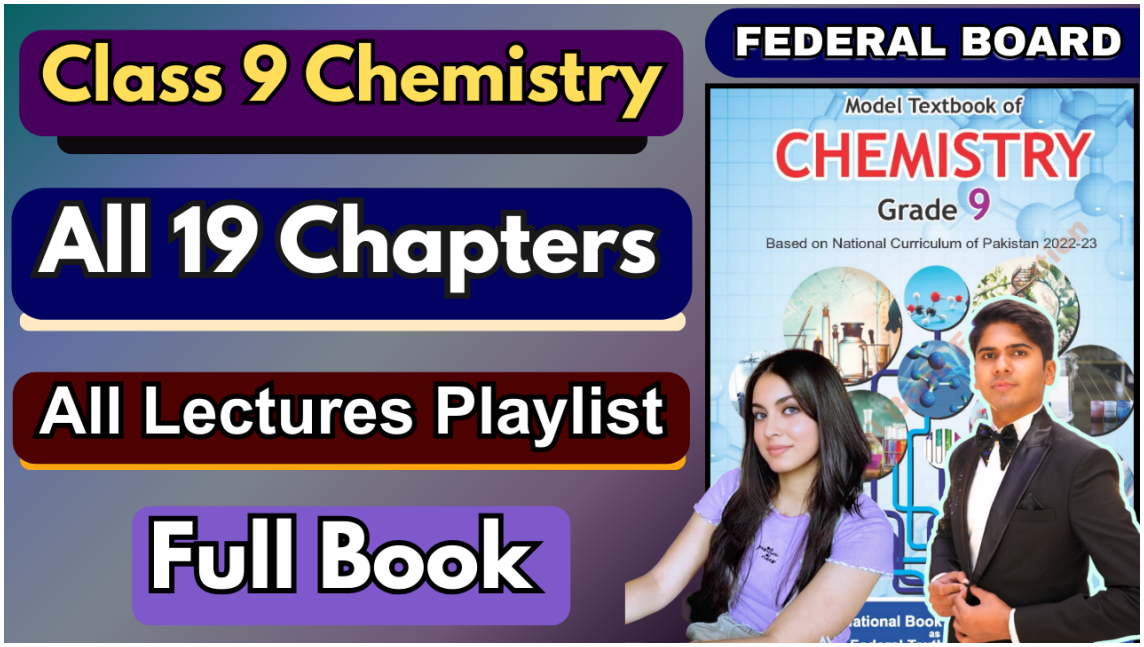


The study of chemistry revolves around the study of matter which is all around us; not only is the entire world made up of matter but so are we, so are the objects that we use. From this we can derive the definition of matter: Anything that has mass and occupies space is called matter.

## 2.1. State of Matter

Matter is anything that **has mass and occupies space**. It is composed of atoms, which are the fundamental building blocks of all substances. The different materials around us appear different because atoms in matter are arranged differently.

### Four States of Matter

- Solid
- Liquid
- Gas
- Plasma

### Why States of Matter Differ

States differ due to **arrangement and movement of particles** and the **strength of intermolecular forces**. **Energy** affects the state: When heated $\rightarrow$ solids melt, liquids evaporate, gases ionize into plasma. When cooled $\rightarrow$ gases condense, liquids freeze.

### Additional/Modern States

- **Liquid Crystal:** Formed when some crystalline solids melt into cloudy liquids that have properties of both solids and liquids. Exists in a limited temperature range.
- **Bose–Einstein Condensates (BEC):** Formed when atoms are cooled very close to **absolute zero**. Found in superfluids and superconductors.

### Macroscopic Properties of Matter (Visible to Naked Eye)

- Density
- Fluidity
- Compressibility

### Properties of States of Matter (From Table 2.1)

- **Property - Gas**
  - **Density:** Low
  - **Volume:** Variable (fills container)
  - **Shape:** Variable (takes shape of container)
  - **Intermolecular Force:** Very weak
  - **Compressibility:** High
  - **Fluidity:** Yes (flows freely)
- **Property - Liquid**
  - **Density:** High
  - **Volume:** Fixed
  - **Shape:** Variable (takes shape of container)
  - **Intermolecular Force:** Medium
  - **Compressibility:** Low
  - **Fluidity:** Yes (flows freely)
- **Property - Solid**
  - **Density:** Highest
  - **Volume:** Fixed
  - **Shape:** Fixed
  - **Intermolecular Force:** Strongest
  - **Compressibility:** Very low
  - **Fluidity:** No

## 2.2. Changes in State of Matter

Matter can change its state by changing temperature or pressure.

### Processes of State Change

- **Melting (Fusion):** Solid to Liquid. Happens at the melting point.
- **Freezing (Solidification):** Liquid to Solid. Happens at the freezing point.
- **Evaporation (Vaporization):** Liquid to Gas. Occurs below the boiling point.
- **Boiling:** Liquid to Gas. Occurs at the boiling point (100°C for water).
- **Condensation:** Gas to Liquid.
- **Sublimation:** Solid to Gas directly (e.g., dry ice) or Gas to Solid directly (deposition).

### Factors Affecting State Change

- **Kinetic Energy:** Higher kinetic energy of particles (due to heat) breaks intermolecular forces, leading to a change of state.
- **Temperature and Pressure:** These factors determine the state of matter.

## 2.3. Classification of Matter

Matter can be classified into two main categories: **Pure Substances** and **Mixtures**.

### 2.3.1. Pure Substances

Pure substances are composed of only one type of particle (atom or molecule) and have fixed, distinct chemical and physical properties (e.g., fixed melting point). They are divided into Elements and Compounds.

**a) Element**

- Contains only **one type of atom**.
- Cannot be broken down into simpler substances by ordinary chemical means.
- Examples: Hydrogen (H), Oxygen (O), Iron (Fe).

**b) Compound**

- Formed when **two or more different elements** combine chemically in a **fixed ratio**.
- The properties of the compound are completely different from its constituent elements.
- Can be separated by chemical methods (e.g., electrolysis).
- Examples: Water ($H_2O$), Salt (NaCl), Sugar ($C_{12}H_{22}O_{11}$).

### 2.3.2. Mixtures

A mixture is formed when **two or more substances** (elements or compounds) are combined physically in a **variable ratio**. They retain their original properties and can be separated by physical methods.

**a) Homogeneous Mixture**

- Has a **uniform composition** throughout.
- Only **one phase** is visible.
- Also called a **solution**.
- Examples: Sugar dissolved in water, air, alloys.

**b) Heterogeneous Mixture**

- Has a **non-uniform composition**.
- The components are **not evenly distributed** and are easily distinguishable.
- **More than one phase** is visible.
- Examples: Sand in water, oil and water.

## 2.4. Solutions

A solution is a **homogeneous mixture** of two or more substances. It is made of a solute and a solvent.

### 2.4.1. Solute

The component that is **dissolved** (usually the lesser quantity).

### 2.4.2. Solvent

The component that **dissolves the solute** (usually the greater quantity).

### 2.4.3. Types of Solutions

Solutions exist in all three physical states: solid, liquid, and gas.

- **Gas in Gas:** Air (Oxygen in Nitrogen)
- **Gas in Liquid:** Soda water (Carbon dioxide in water)
- **Liquid in Liquid:** Alcohol in water
- **Solid in Liquid:** Sugar in water
- **Gas in Solid:** Hydrogen gas absorbed in Palladium metal
- **Liquid in Solid:** Mercury in Copper (amalgam)
- **Solid in Solid:** Alloys like Brass (Zinc in Copper)

### 2.4.4. Solubility and The Term ”Aqueous”

**Aqueous Solution:** A solution where **water is the solvent** (e.g., salt in water).

### 2.4.5. Saturated, Unsaturated, and Supersaturated Solutions

- **Unsaturated solution:** Can dissolve **more solute** at a given temperature.
- **Saturated solution:** Cannot dissolve any **more solute** at a given temperature.
- **Supersaturated solution:** Contains **more solute** than is normally possible to dissolve (unstable: crystals form when a seed crystal is added).

### 2.4.6. Concentrated and Dilute Solutions

- **Dilute solution:** contains a **small amount** of solute.
- **Concentrated solution:** contains a **large amount** of solute.
- Example: Brine (concentrated NaCl solution)

Adding more solvent $\rightarrow$ decreases concentration.

### 2.4.7. Solubility

Solubility = maximum amount of solute that dissolves in a **specific amount of solvent** at a **specific temperature**. Depends on:

- Nature of solute
- Nature of solvent
- Temperature
- Pressure

### 2.4.8. Effect of Temperature on Solubility

Temperature affects solubility differently for different solutes:

- **a) Solubility increases with temperature** Examples:
  - KCl
  - $NH_4Cl$
- **b) Solubility decreases with temperature** Examples:
  - $Na_2SO_4$
  - $Ca(OH)_2$

Solubility curves show how solubility varies with temperature.

## 2.5. Colloids & Suspensions

### Colloid

A **heterogeneous mixture** with particle sizes between **1–1000 nm**. Characteristics:

- Particles do **not settle** on standing.
- Particles scatter light $\rightarrow$ **Tyndall effect** (path of light becomes visible).
- Called **false solutions** or **colloidal dispersions**.

Examples: Milk, starch, blood, jelly, ink, toothpaste.

### Suspension

A **heterogeneous mixture** in which **solid particles remain dispersed** in a liquid but do not dissolve. Characteristics:

- Particle size **> 1000 nm**.
- Particles **settle on standing**.
- Visible to naked eye.

Examples: Chalk in water Milk of magnesia Paints


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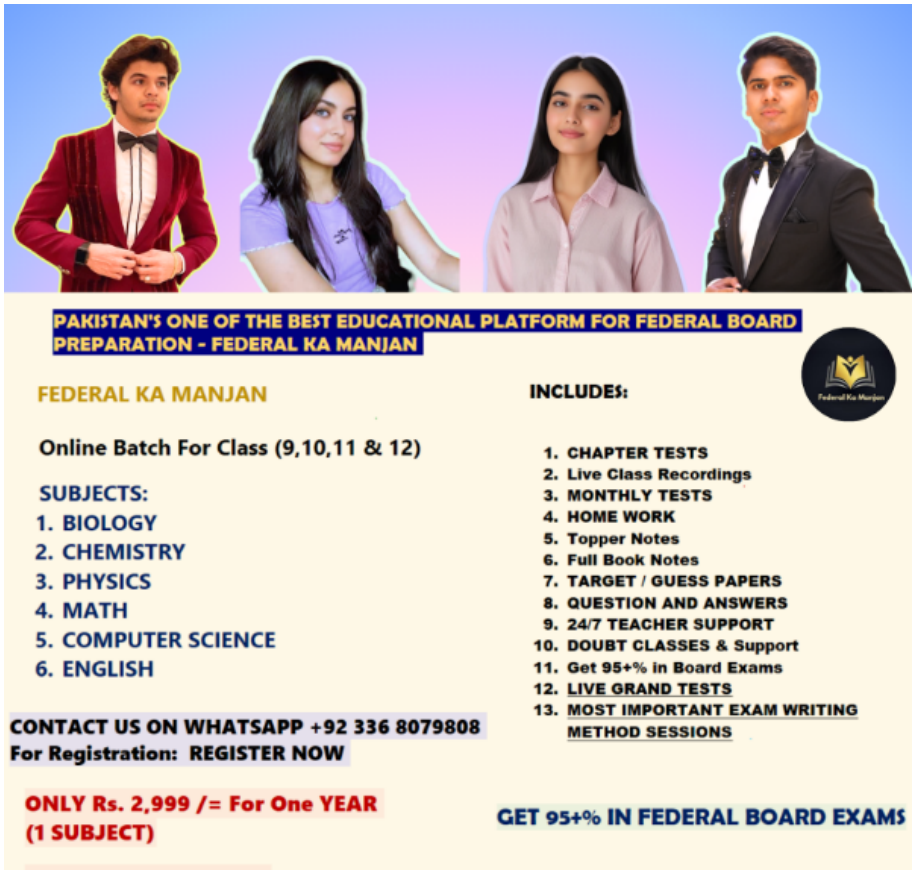
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