Unit 3

PHYSICAL STATES OF MATTER

Introduction to the States of Matter

Matter is anything that occupies space and has mass. It is the material of the universe and exists primarily in three states: solid, liquid, and gas. For example, water can exist in these three states:

- **Steam (water vapor):** Gaseous state of water
- **Water:** Liquid state of water
- **Ice:** Solid state of water

Changes in the state of matter, such as ice melting, water freezing, boiling, or steam condensing, are common in everyday life. These changes occur due to variations in temperature and/or pressure.

States of Matter:

1. **Solid:**
   - **Characteristics:** Solids have a fixed shape and volume. They are rigid and cannot be compressed. The particles in a solid are tightly packed and organized, vibrating at fixed points. Examples include metals, chalk, and sand.
   - **Density:** Solids have higher densities than liquids because their particles are closer together.
2. **Liquid:**
   - **Characteristics:** Liquids have a definite volume but no definite shape, taking the shape of their container. They are slightly compressible and can flow. The particles in a liquid vibrate and shift positions. Examples include water and oil.
   - **Density:** Liquids are denser than gases but less dense than solids due to the closer arrangement of particles compared to gases.
3. **Gas:**
   - **Characteristics:** Gases have neither a definite volume nor shape, taking the shape and volume of their container. Gases are highly compressible because of the large spaces between particles. Examples include air and hydrogen.
   - **Density:** Gases have the lowest density because their particles are far apart.

4. **Plasma:**
    - **Characteristics:** Plasma is a state of matter at extremely high temperatures where atoms lose electrons and form ionized gases. Plasmas have no fixed shape or volume and are less dense than solids or liquids. Examples include the sun and other stars.

Kinetic Theory of Matter: The kinetic theory explains the behavior of particles in different states of matter:

1. **Particles in Motion:** All matter is composed of particles (ions, atoms, molecules) in constant motion.
2. **Energy Levels:** Particles have kinetic energy (energy of movement) and potential energy. Solids have the least, while gases have the most.
3. **State Changes:** The state of matter depends on the energy and motion of the particles. Changes in state (like melting or boiling) occur with changes in energy.
4. **Temperature:** The temperature measures the average kinetic energy of the particles.
5. **Particle Spacing:** There is space between particles, which increases from solids to gases.
6. **Intermolecular Forces:** Attractive forces between particles are stronger in solids and weaker in gases.

Properties of Matter Explained by Kinetic Theory:

1. **Gases:**
    - No fixed shape or volume.
    - Highly compressible due to large spaces between particles.
    - Low density compared to solids and liquids.
    - Exert pressure in all directions.
    - Easily flow and diffuse.
2. **Liquids:**
    - Definite volume but no definite shape.
    - Higher density than gases due to closer particle arrangement.
    - Slightly compressible.
    - Flow easily, but slower than gases.
3. **Solids:**
    - Definite shape and volume.
    - Higher density than liquids and gases.
    - Difficult to compress due to strong particle attraction.
    - Not fluid, do not flow easily.

Activity :

1. **Substance Comparison:**
   - **Solid Example:** Ice
     - **Motion of Particles:** Vibrate in place.
     - **Distance Between Particles:** Very close.
     - **Attraction Between Particles:** Strong.
   - **Liquid Example:** Water
     - **Motion of Particles:** Move around but stay close.
     - **Distance Between Particles:** Moderately close.
     - **Attraction Between Particles:** Moderate.
   - **Gas Example:** Air
     - **Motion of Particles:** Move freely and rapidly.
     - **Distance Between Particles:** Very far apart.
     - **Attraction Between Particles:** Weak.
2. **Highest Kinetic Energy:**
   - The particles in a **gas** possess the highest kinetic energy.

Kinetic Molecular Theory of Gases:

1. Gases consist of particles in constant, straight-line motion.
2. Particles do not attract or repel each other; collisions are elastic.
3. Large spaces exist between gas particles, making gases easily compressible.
4. Average kinetic energy of gas particles is proportional to temperature.

Understanding these principles helps explain the behavior of gases and their interaction with temperature, pressure, and volume.

The Gas Laws

The gas laws describe the relationships between the volume, pressure, temperature, and quantity of a gas. They have evolved from centuries of research and experimentation. Here is an overview of the key gas laws and related concepts:

1. Pressure

Definition: Pressure is the force exerted per unit area by gas particles colliding with the surfaces around them.

Formula: Pressure=$\frac{Force}{\text{Area}}$

Units:

- **Atmosphere (atm)**
- **Pascal (Pa)**: 1 Pa=1 N/m^2
- **Torr**
- **Millimeter of Mercury (mmHg)**

Conversions:

- 1 atm=760 mmHg=101,325 Pa

2. Volume

Definition: Volume is the space occupied by a gas.

Units:

- **Cubic meter (m^3)**
- **Cubic centimeter (cm^3)**: 1 $cm^3=1\times10^{-6}$ m^3
- **Liter (L)**: 1 L=1 dm^3=1000 cm^3

3. Temperature

Definition: Temperature measures the degree of hotness or coldness.

Scales:

- **Celsius (°C)**
- **Fahrenheit (°F)**
- **Kelvin (K)**: Used in gas law calculations.
  - **Conversion:** K=°C+273

4. Boyle's Law

Definition: At constant temperature, the volume of a fixed amount of gas is inversely proportional to its pressure.

Mathematical Expression: $V\propto\frac{1}{P}$

PV=k Where k is a constant.

Example: If the pressure doubles, the volume halves. For two states of the gas:

$P_1V_1=P_2V_2$

5. Charles' Law

Definition: At constant pressure, the volume of a gas is directly proportional to its Kelvin temperature.

Mathematical Expression:

$V \propto T$

$\frac{V1}{\mathrm{T1}} = \frac{V2}{\mathrm{T2}}$

Example: If the temperature increases, the volume increases proportionally, provided the pressure is constant.

6. Gay-Lussac's Law

Definition: At constant volume, the pressure of a gas is directly proportional to its Kelvin temperature.

Mathematical Expression:

$P \propto T$

$\frac{P1}{\mathrm{T1}} = \frac{P2}{\mathrm{T2}}$

Example: If the temperature of a gas increases, the pressure increases proportionally if the volume is held constant.

7. The Combined Gas Law

Definition: Combines Boyle's, Charles', and Gay-Lussac's laws to relate pressure, volume, and temperature when all three are changing.

Mathematical Expression: $\frac{PV}{\mathrm{T}} = k$

$\frac{P1V1}{\mathrm{T1}} = \frac{P2V2}{\mathrm{T2}}$

8. Avogadro's Law

Definition: At the same temperature and pressure, equal volumes of gases contain equal numbers of moles.

Mathematical Expression:

$V \propto n$

$\frac{V1}{n1} = \frac{V2}{n2}$

Ideal Gas Equation

The **ideal gas equation** is a fundamental concept in chemistry that describes the behavior of gases under various conditions. It combines three basic gas laws into one equation:

1. **Boyle's Law:** At constant temperature and number of moles, the volume of a gas is inversely proportional to its pressure.

   $V \propto \frac{1}{P}$

2. **Charles' Law:** At constant pressure and number of moles, the volume of a gas is directly proportional to its temperature in Kelvin.

   $V \propto T$

3. **Avogadro's Law:** At constant pressure and temperature, the volume of a gas is directly proportional to the number of moles of gas.

   $V \propto n$

Combining these laws, we get:

$V \propto \frac{nT}{P} V$

To turn this proportionality into an equation, we introduce a proportionality constant, R, known as the **gas constant**:

$V = \frac{nRT}{P}$

Rearranging gives us the **ideal gas equation**:

$PV = nRT$

Where:

- P = pressure of the gas

- V = volume of the gas
- n = number of moles of the gas
- R = ideal gas constant (0.0821 L·atm/mol·K)
- T = temperature in Kelvin

At Standard Temperature and Pressure (STP):

- n =1 mol
- V=22.4 L
- T=273 K
- P =1 atm

Using these values in the ideal gas equation, we can calculate R:

R=$\frac{PV}{\mathrm{nT}}$ =$\frac{1\ atm\times 22.4\ L}{1\ \mathrm{mol}\times 273\ \mathrm{K}}$ =0.0821 Latm/molK

The ideal gas equation is useful for calculating any one of the gas properties (P, V, T, or n) when the other three are known.

Graham's Law of Diffusion

Graham's Law of Diffusion describes how gases diffuse (spread) and effuse (pass through a small hole) based on their molecular masses.

Key Concepts:

- **Diffusion:** The process by which gas molecules spread throughout a container.
- **Effusion:** The process by which gas molecules pass through a tiny opening.

Thomas Graham discovered that lighter gases (with lower molecular mass) diffuse faster than heavier gases. This is quantified as:

$r \propto \frac{1}{\sqrt{M}}$ Where:

- r = rate of diffusion
- M = molar mass of the gas

For two gases (Gas 1 and Gas 2), Graham's law can be expressed as:

$$\frac{r1}{\mathrm{r2}} = \sqrt{\frac{M2}{\mathrm{M1}}}$$

Where:

- r_1 and r_2 are the rates of diffusion of Gas 1 and Gas 2, respectively.
- M_1 and M_2 are the molar masses of Gas 1 and Gas 2, respectively.

Example Calculation:

If methane (CH_4) diffuses twice as fast as an unknown gas, the relationship can be written as:

$r_{CH4}=2r_x$

Substitute $r_{CH4}=2r_x$ into Graham's Law:

$$\frac{2rx}{\mathrm{rx}} = \sqrt{\frac{Mx}{\mathrm{MCH4}}}$$

Thus:

$$2=\sqrt{\frac{Mx}{16}}$$

Square both sides to solve for M_x:

$$4 = \frac{Mx}{16}$$

M_x=64 g/mol

So, the molecular mass of the unknown gas is 64 g/mol.

Rate of Diffusion and Time: The rate of diffusion is inversely proportional to the time taken to diffuse. For two gases, the time taken is:

$$\frac{t1}{\mathrm{t2}} = \sqrt{\frac{\mathrm{M2}}{\mathrm{M1}}}$$

Where t_1 and t_2 are the times for gases 1 and 2 to diffuse, respectively.

The Solid State

Freezing and Melting Points

- **Freezing Point:** The temperature at which a liquid turns into a crystalline solid (freezes). This temperature is the same as the melting point. At this temperature, the solid and liquid phases are in dynamic equilibrium.
- **Melting Point:** The temperature at which a solid changes into a liquid. During melting, the ordered structure of the solid breaks down, allowing particles to move freely and form a liquid.

Processes

- **Melting (Fusion):** The process where a solid absorbs heat and changes into a liquid. This process requires energy and is endothermic. For example, ice melts at 0°C.
- **Freezing (Solidification):** The process where a liquid loses heat and becomes a solid. It involves the formation of a regular pattern as particles come closer and is the reverse of melting.

Heat of Fusion

- **Heat of Fusion (ΔH_{fus}):** The amount of heat required to convert one mole of a solid into a liquid at its melting point. For example, the molar heat of fusion for ice is 6.01 kJ/mol at 0°C. This is the energy required to overcome the attractive forces in the solid.
- **Heat of Solidification (ΔH_{cryst}):** The energy released when one mole of a liquid solidifies. It is equal in magnitude but opposite in sign to the heat of fusion:

  $\Delta H_{cryst} = -\Delta H_{fus}$.

Sublimation and Deposition

- **Sublimation:** The process where a solid changes directly into a gas without passing through the liquid phase. This process requires energy and is endothermic. The enthalpy of sublimation (ΔH_{sub}) is the energy needed for this transition.
- **Deposition:** The reverse of sublimation, where a gas changes directly into a solid. This process releases energy and is exothermic. The enthalpy of deposition (ΔH_{dep}) is equal in magnitude but opposite in sign to the enthalpy of sublimation:

  $\Delta H_{sub} = \Delta H_{fus} + \Delta H_{vap}$ and $\Delta H_{sub} = -\Delta H_{dep}$.

Heating Curve

- **Heating Curve:** A plot of temperature versus heat added, showing the relationship between phase changes and enthalpy. The curve typically includes regions where temperature rises with added heat and plateaus where temperature remains constant during phase transitions.

Key Points

- **Phase Change Observations:**
  - Melting and freezing occur at the same temperature but are opposite processes.
  - Sublimation and deposition involve transitions directly between solid and gas phases, bypassing the liquid phase.
- **Energy Changes:**
  - **Endothermic Processes:** Melting, vaporization, and sublimation (energy is absorbed).
  - **Exothermic Processes:** Freezing, condensation, and deposition (energy is released).