Unit 3 : Physical States of Matter

- You're likely already familiar with the three common states of matter – solid, liquid, and gas. **The key difference between these states lies in the arrangement and motion of the particles that compose them.**
- **Solids** have a fixed shape and volume. This is because the particles in a solid are tightly packed together in a highly ordered arrangement and have very limited freedom of movement – they can only vibrate about their fixed positions. This strong attraction and limited movement result in high density and low compressibility.
- **Liquids** have a fixed volume but assume the shape of their container. The particles in a liquid are still relatively close together but have more freedom of movement than in a solid – they can move past one another and rotate. This explains why liquids flow and can be poured. While liquids are denser than gases, they are less dense than solids, and while more compressible than solids, they are less compressible than gases.
- **Gases** have no fixed shape or volume and expand to fill their container completely. The particles in a gas are much farther apart than in solids or liquids and are in constant, random motion, moving at high speeds in straight lines until they collide with each other or with the walls of their container. These collisions with the container walls are what we experience as gas pressure. Gases have low density and high compressibility.
- Beyond these three familiar states, there's also **plasma**, often considered the "fourth state of matter." Plasma forms at extremely high temperatures when atoms gain enough energy for some or all of their electrons to be stripped away, creating a mixture of ionized particles, neutral atoms, and electrons. Plasma differs from a gas due to the presence of charged particles, which makes it electrically conductive and responsive to magnetic fields.
- **The Kinetic Molecular Theory of Matter provides a model for understanding these differences in the states of matter.** This theory is based on several key postulates:
    1. Matter is composed of tiny particles (atoms, molecules, or ions).
    2. These particles are in constant, random motion.
    3. The spaces between the particles are large compared to the size of the particles themselves.
    4. The average kinetic energy of the particles is proportional to the absolute temperature of the system.
    5. There are attractive forces between the particles, and the strength of these forces depends on the distance between the particles (stronger at closer distances).
- **The kinetic molecular theory successfully explains many observed properties of matter in different states.** For example:

    - The particles in gases have a high degree of freedom of movement and weak attractive forces, allowing them to expand to fill their containers.
    - The particles in liquids are more closely spaced and experience stronger attractive forces than in gases, giving liquids a fixed volume and surface tension.

learnify

    - The strong attractive forces and limited movement of particles in solids explain their fixed shape and volume.
- **Unit 3 delves into a detailed explanation of the gaseous state, including the gas laws that describe the relationships between pressure, volume, temperature, and the amount of gas.** These laws, derived from experimental observations, provide a mathematical framework for understanding and predicting gas behaviour:
    - **Boyle's Law** states that at a constant temperature, the volume of a fixed mass of gas is inversely proportional to its pressure. In other words, as pressure increases, volume decreases, and vice-versa. Mathematically: $P_1V_1 = P_2V_2$.
    - **Charles's Law** states that at constant pressure, the volume of a fixed mass of gas is directly proportional to its absolute temperature (in Kelvin). This means that as temperature increases, volume also increases, and vice-versa. Mathematically: $V_1/T_1 = V_2/T_2$.
    - **Gay-Lussac's Law** states that at a constant volume, the pressure of a fixed mass of gas is directly proportional to its absolute temperature (in Kelvin). Therefore, as temperature increases, pressure increases, and vice-versa. Mathematically: $P_1/T_1 = P_2/T_2$.
    - **Avogadro's Law** states that at constant temperature and pressure, the volume of a gas is directly proportional to the number of moles (or molecules) of gas present. In other words, equal volumes of different gases at the same temperature and pressure contain the same number of molecules. Mathematically: $V_1/n_1 = V_2/n_2$.
- **These gas laws can be combined into a single equation known as the Ideal Gas Law:** $PV = nRT$.
    - P = Pressure (usually in atmospheres or kilopascals)
    - V = Volume (usually in litres or cubic metres)
    - n = Number of moles of gas
    - R = The ideal gas constant (0.0821 L atm/mol K or 8.314 J/mol K)
    - T = Temperature (always in Kelvin)
- **The ideal gas law is a powerful tool for predicting the behaviour of gases, especially under conditions of high temperature and low pressure, where real gases most closely approximate ideal behaviour.**
- **important concepts related to phase transitions:**
    - **Melting** is the process by which a solid changes to a liquid.
    - **Freezing** is the reverse process of a liquid changing to a solid.
    - **Boiling** is the process of a liquid changing to a gas (vapourisation) throughout the bulk of the liquid.
    - **Condensation** is the reverse process of a gas changing to a liquid.
    - **Sublimation** is the process of a solid changing directly to a gas without passing through the liquid phase.
    - **Deposition** is the reverse process of a gas changing directly to a solid.
- **Each of these phase transitions is associated with an energy change (enthalpy change), which can be either endothermic (heat is absorbed) or exothermic (heat is released).** For example, melting is endothermic, while freezing is exothermic.

Learnify