SLO 6: States of Matter

6.1 Kinetic Molecular Theory and Phase Changes

6.1.1 Compare The Physical States Of Matter Based On The Intermolecular Forces Present Between Their Molecules

Solids:
Possess the strongest intermolecular forces. These forces are strong enough to lock the molecules into a fixed, rigid, and orderly arrangement. Molecules can only vibrate about their fixed positions, resulting in a definite shape and volume. Solids are virtually incompressible due to the minimal space between particles.

Liquids:
Have moderate intermolecular forces, which are weaker than those in solids but stronger than those in gases. These forces are strong enough to keep the molecules in contact with each other, resulting in a definite volume and low compressibility, but weak enough to allow the molecules to move past one another. This movement allows liquids to flow and take the shape of their container.

Gases:
Exhibit the weakest or practically negligible intermolecular forces. The kinetic energy of the molecules is much greater than the attractive forces between them, allowing the molecules to move freely, rapidly, and independently in all directions. This results in no fixed shape or volume, and gases expand to fill their entire container and are highly compressible.

Feature	Solid	Liquid	Gas
Intermolecular forces	Very Strong	Moderate	Very Weak / Negligible
Particle arrangement	Tightly packed, ordered, fixed positions	Close together, no regular arrangement, can slide past each other	Far apart, random arrangement, free-flowing
Molecular motion	Vibration about fixed positions only	Random movement and flow	Rapid, random, and independent movement
Shape	Definite shape	takes shape of container	takes shape of container
Volume	Definite volume	Definite volume	expands to fill container
Compressibility	Almost incompressible	Slightly compressible	Highly compressible

6.1.2 Explain Phase Changes Due To Changes In Temperature And Pressure Affecting The Arrangement And Motion Of Particles Within A Substance

Phase changes are physical processes where a substance transitions between solid, liquid, and gas states due to changes in external conditions like temperature and pressure. These changes occur as the balance between the particles' kinetic energy and the intermolecular forces (IMFs) is altered, which in turn affects the arrangement and motion of the particles.

Effect of Temperature

- Increased Temperature: Increases the kinetic energy of particles, allowing them to overcome IMFs and move further apart into a less ordered, more energetic state.
- Decreased Temperature: Decreases the kinetic energy of particles, allowing IMFs to pull them closer together into a more ordered, less energetic state.

Effect of Pressure

- Increased Pressure: Forces particles closer together, increasing the effectiveness of intermolecular forces and favoring states with higher density (liquid or solid).
- Decreased Pressure: Allows particles more space to move, favoring the gaseous state. Water boils at a lower temperature at higher altitudes due to lower atmospheric pressure.

Phase Change	Transition	Energy Change	Particle Motion & Arrangement
Melting	Solid → Liquid	Endothermic (Energy Gain)	Particles vibrate more vigorously until they break free from their fixed positions and can flow past each other.
Freezing	Liquid → Solid	Exothermic (Energy Loss)	Particles slow down, and IMFs pull them into a rigid, fixed, crystalline lattice structure.
Vaporization	Liquid → Gas	Endothermic (Energy Gain)	Particles gain enough energy to completely overcome all IMFs, becoming widely separated and moving randomly and rapidly.
Condensation	Gas → Liquid	Exothermic (Energy Loss)	Particles lose energy, slow down, and IMFs draw them closer together into contact, although they still flow.
Sublimation	Solid → Gas	Endothermic (Energy Gain)	Particles directly gain enough energy to skip the liquid phase entirely, transitioning straight to a widely separated gaseous state.
Deposition	Gas → Solid	Exothermic (Energy Loss)	Particles directly lose significant energy, skipping the liquid phase and locking into a fixed, rigid solid structure.

6.1.3 Explain Melting And Boiling Point Of Substances As A Criterion To Check Purity

Melting and boiling points of a substance are used to check for purity because pure substances have sharp (specific) melting and boiling points, whereas impure substances melt and boil over a range of temperatures.

Principle of Purity Testing

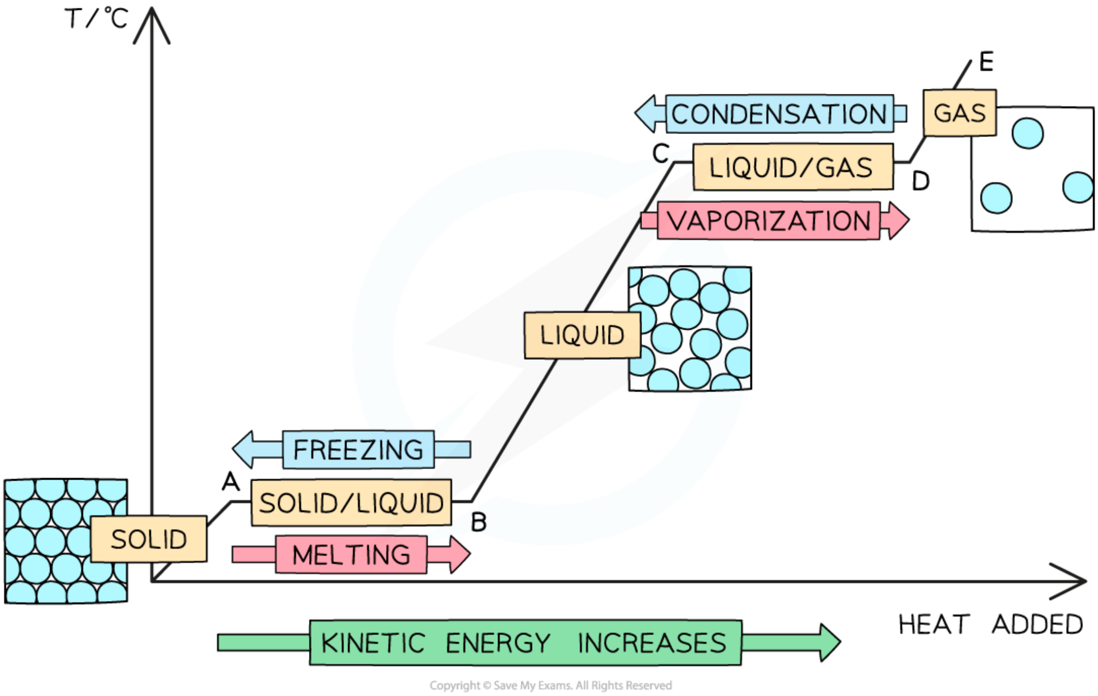
Every pure element or compound has a unique and fixed temperature at which it transitions from solid to liquid (melting point) or from liquid to gas (boiling point) at a given pressure. This is because a pure substance consists of only one type of particle, and intermolecular forces between all particles are uniform. Therefore, a consistent, specific amount of energy is required to overcome these forces simultaneously during a phase change, resulting in a single, sharp transition temperature.

Effect of Impurities

1. *Melting Point Depression and Broadening:* Impurities generally lower the melting point and cause the substance to melt over a wider range of temperatures, rather than at a single point.
2. *Boiling Point Elevation and Broadening:* Impurities generally raise the boiling point of a liquid and cause it to boil over a wider range of temperatures. This is because the impurities lower the vapor pressure of the liquid, requiring a higher temperature to reach the point where the vapor pressure equals the surrounding atmospheric pressure.

6.1.4 interpret heating and cooling curves in terms of kinetic theory

Heating and cooling curves visually represent the relationship between energy added/removed over time and the resulting temperature changes and phase transitions of a substance. The kinetic theory of matter explains these phenomena by distinguishing between changes in temperature (K.E) and changes in intermolecular forces (P.E) of the particles.



Heating

- When a solid is heated, the particles gain kinetic energy and vibrate faster. At melting point, the energy transferred to the substance is now used to overcome the intermolecular forces of attraction holding the particles in their solid structure. Here, the temperature of the substance stops increasing. This is melting and will continue until all the particles have reached the liquid state.
- As substance is continually heated, the liquid particles gain more kinetic energy and move over one another faster. At boiling point, the energy transferred to the substance is now used to overcome the intermolecular forces of attraction holding the particles in their liquid structure. The temperature of the substance stops increasing. This is evaporation or vaporization and will continue until all the particles have reached the gaseous state.

Cooling

- The process is repeated backwards for cooling as energy is transferred away. A gas turns back into liquid through condensation and liquid turns back into a solid through freezing.
- When energy is transferred away from a gas, the temperature starts to decrease as the particles lose kinetic energy and move slower. When the temperature of the substance reaches its boiling point, the energy transferred away from the substance is now taken from the energy used to overcome the intermolecular forces of attraction.
- The temperature of the substance stops decreasing and particles now succumb to those intermolecular forces of attraction and are held together in their liquid structure. This is condensing and will continue until all the particles have reached the liquid state.
- When energy is transferred away from a liquid, the temperature starts to decrease as the particles lose more kinetic energy and move even slower.
- When the temperature of the substance reaches its melting point, the energy transferred away from the substance is now taken from the energy used to overcome the intermolecular forces of attraction.
- The temperature of the substance stops decreasing. The particles now succumb to those intermolecular forces of attraction and are held together in their solid structure. This is freezing and will continue until all the particles have reached the solid state.

6.2 Gaseous State

6.2.1 Explain The Following Properties Of Gases In Terms Of Kinetic Theory

1. Diffusion: Gases can diffuse very rapidly. Diffusion is defined as spontaneous mixing up of molecules by random motion and collisions to form a homogeneous mixture. Rate of diffusion depends upon the molecular mass of the gases. Lighter gases diffuse rapidly than heavier ones. For example, H_2 diffuses four times faster than O_2 gas.
2. Effusion: Effusion is defined as the escaping of gas molecules through a tiny hole into a space with lesser pressure. For example, when a tire gets punctured, air effuses out.
3. Density: Density is defined as mass per unit volume. Gases have low density than liquids and solids due to light mass and more volume occupied by the gas molecules. Gas density is expressed in g/dm^3. The density of gases increases by cooling because their volume decreases. For example, at normal atmospheric pressure, the density of oxygen gas is 1.4 g/dm^3 at 20°C and 1.5 g/dm^3 at 0°C.
4. Compressibility: Compressibility is defined as the ability of gases to reduce their volume under pressure. Gas has mostly empty space between its molecules. Applying pressure forces molecules closer, reducing the volume significantly. For example, LPG gas is compressed into a small cylinder by reducing its volume 200–300 times.

6.2.2 Relate Qualitatively The Effect Of The Following Factors To The Rate Of Diffusion

Molecular Mass

According to the kinetic theory, all particles of different substances at the same temperature possess the same average kinetic energy. To maintain equal kinetic energy, a particle with a larger mass must move with a lower average velocity, while a particle with a smaller mass must move with a higher average velocity. The faster a particle moves, the quicker it spreads out and mixes with other particles. A quantitative relationship is provided by Graham's Law of Diffusion, which states the rate of diffusion is inversely proportional to the square root of its molecular mass.

$$diffusion \propto \frac{1}{\sqrt{molecular\ mass}}$$

Temperature

An increase in temperature adds thermal energy to the system, which increases the average kinetic energy of all particles. As the particles gain more energy, they move with greater average velocities, leading to more frequent and faster collisions and a more rapid spread throughout the medium. Conversely, when temperature decreases, particles slow down, and diffusion becomes a slower process.

$$temperature \propto diffusion$$

6.3 Laws Related to Gases

6.3.1 Relate Changes In Pressure And Volume Of A Gas Using Boyle’s Law

Boyle’s Law describes the quantitative relationship between the pressure and volume of a fixed mass of gas at constant temperature. It states that the pressure of an ideal gas is inversely proportional to its volume.

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

$$P_i V_i \propto P_f V_f$$

- Decreasing Volume (Increasing Pressure): When the volume of the container is reduced (e.g., by pushing down a piston), the gas particles are confined to a smaller space. The particles hit the container walls more frequently per unit of time. The increased frequency of collisions results in a higher overall force exerted on the walls, thus increasing the pressure.
- Increasing Volume (Decreasing Pressure): When the volume is increased, the particles have farther to travel between collisions with the container walls. The frequency of collisions decreases, leading to a lower overall force on the walls and a decrease in pressure.

6.3.2 Relate Changes In Temperature And Volume Of A Gas Using Charles's Law

Charles's Law describes the quantitative relationship between the temperature and volume of a fixed amount of gas at constant pressure. It states that the volume of an ideal gas is directly proportional to its absolute temperature (measured in Kelvin).

$$V \propto T$$

$$\frac{V_i}{T_i} \propto \frac{V_f}{T_f}$$

- Increasing Temperature (Increasing Volume): When the temperature of a gas increases, the particles gain more kinetic energy and move faster. These faster-moving particles collide with the container walls more frequently and with greater force. To maintain a constant pressure (as required by Charles's Law), the volume of the container must expand. This expansion increases the surface area of the walls, spreading out the increased force of collisions and keeping the pressure constant.
- Decreasing Temperature (Decreasing Volume): When the temperature decreases, the particles slow down, colliding with the container walls less frequently and with less force. The container volume decreases until the collision frequency is high enough to restore the original constant pressure.

6.3.3 Relate Changes In Number Of Molecules, Volume Of Gas Using Avogadro's Law

Avogadro's Law describes the relationship between the volume of a gas and the number of molecules present, provided the temperature and pressure remain constant. It states that the volume of a gas is directly proportional to the number of moles (amount of substance) of gas present.

$$V \propto n$$

$$\frac{V_i}{n_i} \propto \frac{V_f}{n_f}$$

- Increasing the Number of Molecules (Increasing Volume): When more gas particles are introduced into a system while keeping the original volume the same, there will be more frequent collisions with the container walls, thus increasing the pressure. To maintain a constant pressure (as required by Avogadro's Law), the container's volume must expand to spread out the increased number of collisions over a larger surface area.
- Decreasing the Number of Molecules (Decreasing Volume): Removing gas particles reduces the frequency of collisions with the container walls, which would lower the pressure. To keep the pressure constant, the volume of the container must decrease (contract).

6.4 Liquid State

6.4.1 Explain Properties Of Liquids And The Factors That Affect Them

Vapor Pressure

- *Definition:* Vapor pressure is the pressure exerted by a vapor in thermodynamic equilibrium with its liquid phase at a given temperature in a closed system.
- *Explanation:* In any liquid, some molecules at the surface have enough kinetic energy to escape into the gas (vapor) phase, a process called evaporation. These vapor molecules collide with the container walls and exert pressure.
- *Factors Affecting Vapor Pressure:*
    1. Temperature (Major Factor): Vapor pressure increases significantly as temperature rises. Higher temperatures mean higher average kinetic energy, allowing more molecules to escape the liquid phase and increasing the vapor concentration and collision frequency.
    2. Intermolecular Forces (Major Factor): Stronger IMFs hold molecules tightly within the liquid phase, making it harder for them to escape as vapor. Liquids with strong IMFs (like water) have lower vapor pressure than liquids with weak IMFs (like alcohol or ether).

Boiling Point

- *Definition:* The boiling point is the temperature at which the vapor pressure of a liquid equals the external pressure surrounding the liquid. The liquid turns into a gas rapidly in a bulk process called boiling.
- *Factors Affecting Boiling Point:*
    1. External Pressure (Major Factor): Because boiling occurs when vapor pressure equals external pressure, changes in external pressure directly affect the boiling point. Decreasing external pressure (like at higher altitudes) lowers the boiling point, while increasing pressure raises it.
    2. Intermolecular Forces (Major Factor): Stronger IMFs require significantly higher temperature to overcome and turn the liquid into a gas. Thus, strong IMFs result in a higher boiling point.
    3. Impurities: Impurities generally raise the boiling point.

Freezing Point

- *Definition:* The freezing point is the temperature at which a liquid changes into a solid. At this temperature, the solid and liquid phases exist in equilibrium, and the particles' kinetic energy is low enough for intermolecular forces to lock them into a rigid, ordered structure.
- *Factors Affecting Freezing Point:*
    1. Intermolecular Forces: Stronger IMFs help form the stable solid lattice at higher temperatures, generally leading to higher freezing points.
    2. Pressure: Pressure typically has a minor effect on the freezing point of most liquids. Water is a notable exception; increasing pressure actually lowers its freezing point slightly.
    3. Impurities: Impurities invariably lower the freezing point (freezing point depression), as they disrupt the formation of the pure crystalline structure. This is why salt is used to melt ice on roads.

Density

- *Definition:* Density is a measure of mass per unit volume. In general, liquids are much denser than gases but less dense than solids.
- *Factors Affecting Density:*
    1. Temperature: Density typically decreases as temperature increases because the kinetic energy causes particles to move slightly further apart, increasing the volume.

2. Intermolecular Forces: Stronger IMFs generally result in a more compact structure (lower volume for the same mass) and higher density.
3. Pressure: Pressure has a very minor effect on the density of liquids because they are largely incompressible.

Compressibility

- *Definition:* Compressibility is a measure of how much the volume of a substance decreases when pressure is applied to it.
- *Explanation:* Liquid particles are already very close together, with little empty space between them compared to gases.
- *Factors Affecting Compressibility:*
  1. Intermolecular Forces: Strong IMFs keep particles tightly packed, resisting volume changes when pressure is applied.
  2. Pressure (Minor Effect): Liquids are considered largely incompressible for most practical purposes. Applying high pressure results in only negligible changes in volume. This low compressibility is why hydraulic systems (like car brakes) use liquids effectively to transmit force.

6.4.2 Differentiate Between Evaporation And Boiling

Evaporation	Boiling
Change of liquid into vapor at any temperature below its boiling point.	Rapid change of liquid into vapor at a specific temperature.
Occurs only at the surface of the liquid.	Occurs throughout the entire liquid.
Can happen at any temperature.	Occurs only at or above the boiling point of the liquid.
Relatively slow process.	Relatively fast process.
Water drying from clothes, sweat cooling the body, puddle disappearing on a sunny day	Water boiling in a kettle, pasta cooking in boiling water

6.4.3 Discuss The Effects Of:

Temperature On Vapor Pressure

Vapor pressure increases with temperature. As temperature rises, more liquid molecules gain enough kinetic energy to escape from the liquid surface into the vapor phase. This increases the number of vapor molecules above the liquid, hence higher vapor pressure. The increase is not linear — it is exponential. At higher temperatures, a small increase in temperature causes a very large increase in vapor pressure.

External Pressure On The Boiling Point

For boiling to occur, vapor pressure must equal external pressure. If external pressure is higher (e.g., in a pressure cooker), the liquid must be heated to a higher temperature to generate enough vapor pressure to equal that higher external pressure. Thus, when external pressure increases; boiling point increases.

6.5 Solid State

6.5.1 Explain The Following Physical Properties Of Solids:

1. Melting Point: The temperature at which a solid turns into a liquid at a given pressure (usually 1 atm). It tells us how much thermal energy is needed to break the orderly structure of the solid and allow particles to flow.
2. Density: Mass per unit volume (ρ = mass/volume). It is usually expressed in g/cm^3 or kg/m^3. It tells us how tightly the particles are packed in the solid.

3. Compressibility: The ability of a substance to decrease in volume under external pressure. It is usually measured as bulk modulus (inverse of compressibility). It tells us how much a solid resist being squeezed.

6.5.2 Explain The Following Applications Of Sublimation:

1. Solid Air Fresheners: Solid air fresheners utilize sublimation to release scent gradually over time. The active ingredient (fragrance) is a solid designed to sublime at room temperature. As the solid slowly turns into a gas, the scent molecules diffuse into the surrounding air. The rate of sublimation is typically slow, which ensures the product provides a long-lasting, continuous scent without needing a liquid dispenser or an electrical power source.
2. Dry Ice (solid CO_2): Dry ice is the solid form of carbon dioxide and is one of the most well-known examples of sublimation in everyday life. Carbon dioxide cannot exist as a liquid at standard atmospheric pressure; it sublimes directly from solid to gas at a temperature of about $-78.5°C$. It is an excellent refrigerant because it is extremely cold and leaves no liquid residue ("dry" ice). When dry ice is placed in warm water, it rapidly sublimes and chills the surrounding water vapor, creating a dense, low-hanging fog effect.
3. Mothballs (naphthalene): Traditional mothballs contain naphthalene or paradichlorobenzene, both of which sublime readily at room temperature. The solid crystals slowly turn into vapor without melting first. The vapor is toxic/heavy and fills closets or drawers, repelling or killing clothes moths and their larvae.
4. 3D printing: Special 3D printers print with sublimation inks onto a surface or into a powder bed. After printing, the object is heated (usually 180–200 °C). The solid dyes sublime and penetrate deeply into the plastic (often PLA or special gypsum-based materials), becoming part of the material itself. The result is vibrant, scratch-resistant, full-color 3D prints (used for figurines, prototypes, medical models).

6.6 Types of Solid

6.6.1 Define The Term 'Allotropes'

Allotropes are different structural forms of the same element in the same physical state (usually solid) that have distinctly different physical properties and sometimes different chemical behavior.

6.6.2 Explain Allotropic Forms Of Carbon And Sulphur

Allotrope	Structure	Key Physical Properties	Common Uses
Diamond	3D tetrahedral network	Hardest natural substance, electrical insulator, high melting point, transparent	Jewelry, industrial cutting tools
Graphite	Layered hexagonal sheets	Soft, slippery (layers slide), electrical conductor, opaque	Lubricants, pencils, electrodes
Buckyballs	Spherical	Molecular solid, soluble in organic solvents, unique electronic properties	Nanotechnology, medical research
Coal	Amorphous, irregular	Non-crystalline, variable hardness/density, black, brittle	Fuel for power generation
Rhombic Sulphur	Orthorhombic	Bright yellow crystals, most stable at room temperature	sulfuric acid, vulcanization, fungicides
Monoclinic Sulphur	Monoclinic	Pale yellow needle-like crystals	demonstrate allotropy in chemistry classes
Plastic Sulphur	Non-crystalline	Brown, rubbery, elastic when freshly made	physical changes Sulphur undergoes when superheated and then rapidly cooled