CIE Chemistry A Level

4 : States of Matter

Notes

Gaseous State

Kinetic theory and ideal gases

The assumptions of ideal gases are:

- Molecules are in **constant random motion** in straight lines.
- Molecules are rigid **spheres.**
- **Pressure** is due to molecules **colliding** with the walls of the container.
- All collisions are **elastic** (whether between molecules or between molecules and the walls of the container).
- **Temperature is proportional to the average kinetic energy** of the molecules.

For a gas to approach ideal behaviour:

- Temperature must be high enough above the boiling point so that there are no intermolecular forces between molecules.
- Pressure must be low enough so that the volume of the individual molecules is negligible relative to the volume of the container.

Limitations of ideality:
The ideal gas law has limitations of ideality at high pressures and very low temperatures. This is because at high temperatures the gas molecules get more crowded which isn't accounted for in the equation. At very low temperatures the effect of intermolecular forces is much more prominent as the molecules have less kinetic energy to overcome the attractions.

The ideal gas equation

The ideal gas equation is: **$pV = nRT$**
p - pressure (Pa)
V - volume (m^3)
n - number of moles (mol)
R - gas constant (given in the data book, 8.31441 J K^{-1} mol^{-1})
T - temperature (K)

Converting units:
Pressure: 1 kPa = 1000 Pa
1 atm = 101325 Pa
1 bar = 100 kPa
Volume: 1 m^3 = 1000 dm^3 = 1000000 cm^3
Temperature: 0 K = -273 °C

Molar mass can be determined using the ideal gas equation and the equation **$n = m/M$** where m is mass in g and M is the molar mass.

Liquid State

Liquids contain **randomly arranged particles** which are close together with some gaps. The gaps allow the particles to move. The particles in a liquid have enough energy to prevent the intermolecular forces holding them in a fixed arrangement. Most liquids have a slightly lower density than the solid.

Melting

Heat energy causes the particles in a solid to **vibrate**. Eventually, the particles have enough energy to disrupt the regular arrangement and become a liquid.

Vaporisation

Heat energy causes the particles in a liquid to move fast enough to **break all the forces of attraction** between them and become a gas.

Vapour pressure

When a liquid evaporates in a closed container, the gaseous particles move around above the liquid. When these particles collide with the walls of the container, they exert a pressure called the **vapour pressure**.

Solid State

The lattice structure of different crystalline solids is described below:

- Giant ionic lattice - **positive and negative ions** alternate in a three dimensional structure, held by ionic bonds.
    - Sodium chloride, Na^+ and Cl^- ions.
    - Magnesium oxide, Mg^{2+} and O^{2-} ions.
- Simple molecular (covalent) lattice - molecules held together by **van der Waals forces** in a cubic structure.
    - Iodine, face centred cubic structure.
    - C_{60} has a ball like structure made up of hexagons and pentagons of carbon atoms with van der Waals forces between molecules.
    - Nanotubes are a cylinder of graphene (single sheet of carbon atoms covalently bonded together).
- Giant molecular (covalent) structure.
    - Diamond, each carbon shares an electron with 4 other carbon atoms.
    - Graphite, hexagons of carbon atoms in layers with each carbon atom covalently bonded to 3 carbon atoms, one delocalised electron per carbon atom so van der Waals forces between layers.
    - Graphene, single layer of graphite.
    - Silicon(IV) oxide, similar 3D structure to diamond with oxygen and silicon atoms covalently bonded together.
- Hydrogen-bonded lattice
    - Ice, open lattice structure held by hydrogen bonds between partially positive hydrogen and a lone pair of electrons on oxygen.

- Metallic - positive ions closely packed together with delocalised electrons

Finite resources and recycling

A finite resource is used up faster than it is replaced so it **will run out** if it is continually used. Examples of finite resources are crude oil, copper and aluminium.

Recycling is important to reduce the rate at which resources are used up. It may reduce costs and the environmental impact of materials.

Hydrogen bonding

Hydrogen bonding can affect the physical properties of a substance in the following ways:

- **Boiling and melting point are increased** because hydrogen bonds are **stronger** than van der Waals forces so require more energy to overcome.
- **Viscosity increases** (the liquid becomes more thick and sticky) because the hydrogen bonds hold the molecules together more strongly.
- **Surface tension increases** because hydrogen bonds hold the molecule together more strongly, increasing the amount of force required to break the skin on the surface of the liquid.

Predicting structure and bonding from data

1. A high boiling point indicates a giant structure (ionic, metallic or giant covalent) while a low boiling point indicates simple molecules (or atoms for noble gases).
2. Compounds that are soluble in water tend to be ionic. If the compound is soluble and has a low boiling point, it may be a simple molecule that is small and very polar or able to form hydrogen bonds.
3. If the solid compound conducts electricity, it is likely to be a metal, graphite or graphene. If the substance only conducts electricity when molten or dissolved, it is an ionic compound.
4. The appearance of a substance can be used to distinguish between giant structures. A shiny, malleable and ductile substance is a metal whereas ionic and giant covalent structures tend to be brittle.