CAIE Chemistry A-level

4: States of Matter

Notes

The 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.

Solid State

The lattice structure of different crystalline solids is described below:

- Giant ionic lattice - **positive and negative ions** alternate in a 3D 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.

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.

Bonding and Physical Properties

The physical properties of a substance are affected by the types of bonding it contains:

- **Covalent (giant structures only)** e.g. silicon(IV) oxide, graphite and diamond:
    - High melting points - strong covalent bonds require a large amount of energy to break.
    - Mostly non-conductors - don't contain mobile charged particles (except graphite).
    - Insoluble - covalent bonds in the lattice are too strong to be broken.
- **Ionic** e.g. sodium chloride and magnesium oxide :
    - High melting and boiling points - strong electrostatic attraction between oppositely charged ions requires a lot of energy to break.
    - Electrical conductor - when aqueous or molten, the ions are free to move to conduct electricity. When solid, the ions are in fixed positions so cannot conduct electricity.
    - Soluble in polar solvent - charged parts of the solvent are attracted to the oppositely charged ions.
- **Metallic** e.g. copper:
    - High melting and boiling point - the attraction between the ions and delocalised electrons is strong so a lot of energy is needed to overcome the metallic bonding.
    - Good electrical conductor - contains mobile delocalised electrons which can conduct electricity as a solid.
    - Malleable and ductile - the regular structure and delocalised electrons allow the uniform layers of ions to slide over one another.
- **Hydrogen bonds** e.g. ice:
    - High boiling point - The boiling points are greater than those of molecules with only van der Waals forces between them because hydrogen bonds are stronger.
    - Soluble in water - strong permanent dipoles allow the formation of hydrogen bonds with water.
    - Non-conductors - no mobile charges so are unable to conduct electricity.
- **Van der Waals forces** e.g. iodine and buckminsterfullerene:
    - Low boiling point - forces are weak so require little energy to break. Larger molecules have more van der Waals forces so have higher boiling points.
    - Solubility - unless they react with water, most molecular compounds are insoluble in water because they release too little energy when they dissolve. They are often soluble in organic solvents because they both contain van der Waals forces.
    - Non-conductors - no mobile charges so are unable to conduct electricity.

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.