SECTION 9

GASES AND THE GAS LAWS

The mass of a gas cannot easily be measured by weighing as is done for a liquid or solid. However the volume of the gas's container, and the pressure and temperature of the gas can be measured. This section tells you about experimental laws and a famous hypothesis which allow the determination of the amount of a gas from its volume, pressure and temperature. Then it briefly gives the simple theory that rationalises these laws.

Ideal gas: A gas that obeys **Boyle's law** and **Charles's law**, and hence the **Ideal gas equation**.

Boyle's law: The pressure, p, of a fixed mass of gas is inversely proportional to its volume, V, at constant temperature (i.e. $p \propto 1/V$).

Charles's law: The volume, V, of a fixed mass of gas is proportional to its absolute temperature, T, at constant pressure (i.e. $V \propto T$).

It is found experimentally that at constant pressure and temperature the volume of a gas is directly proportional to the amount of gas, n, and at constant volume and temperature the pressure is directly proportional to the amount of gas, n.

Avogadro's hypothesis: Equal volumes of different gases at the same temperature and pressure contain the same number of molecules (or atoms for monoatomic gases).

All this is expressed by the **ideal gas equation**.

Ideal gas equation: $pV = nRT$ R is called the **universal gas constant**, and has the value 8.314 J K^{-1} mol^{-1}. This equation is particularly useful in determining the amount of a gas from its volume, pressure and temperature.

[e.g. Calculate the mass of dinitrogen in 500 litres at a pressure of 200 kPa and a temperature of 100^{o}C. $n(N_2)$ is obtained from the ideal gas equation, and the mass from the molar mass:

$$n(\mathrm{N_2}) = \frac{pV}{RT} = \frac{200 \times 10^3 \text{ Pa x } 0.500 \text{ m}^3}{8.314 \text{ J K}^{-1} \text{ mol}^{-1} \text{ x } 373 \text{ K}} = 32.2 \text{ mol}$$

$$m(\mathrm{N_2}) = nM = 32.2 \text{ mol x } (2 \text{ x } 14.0 \text{ g mol}^{-1}) = 903 \text{ g }]$$

Ideal gas: a gas for which the individual molecules or atoms occupy negligible volume, and for which there are no attractive or repulsive forces between the molecules (or atoms).

Dalton's law of partial pressures: The total pressure of a mixture of gases is the sum of the **partial pressures** of the individual gases in the mixture. All real gases deviate from ideal behaviour, the deviation increasing with increasing p and decreasing T.

Partial pressure: The pressure that a particular gas in a mixture of gases would exert if it alone occupied the container.

The **kinetic theory of gases** makes the following assumptions for an ideal gas:

1. Gases are made up of molecules whose sizes are negligible compared with the distance between them.
2. There are no forces between the molecules (except in a collision).
3. Between collisions the molecules are constantly moving in straight lines and their motion is completely random.
4. The molecules are constantly colliding with one another and with the walls of the container, all collisions being elastic (i.e. no loss of kinetic energy).
5. The collisions with the walls of the container give rise to the measured gas pressure.

Vapour: An alternative term for gas. It is usually used when it is in contact with the liquid form, solid form or solution of the same substance, or is at a temperature at which it could be made to condense by increasing its pressure. The term **evaporation** illustrates this.

EXERCISES

Assuming the ideal gas equation is obeyed determine:

1. The volume of 50 g of dioxygen at 500 K and 200 kPa

2. The mass of 5 L of HBr at 1000 ^oC and 4 mPa

3. The pressure of 5 kg of argon in a 5 litre cylinder at 25 ^oC

4. The change in pressure if a given volume of xenon at 100 kPa is heated from 298 K to 800 K

5. The total pressure in a 2 litre cylinder containing 2 g of N_2O and 3 g of O_2 at 298 K