Atomic structure

Chemistry A – Atoms, molecules and energy

Properties of solids: not compressible, do not flow, negligible expansion and contraction with changes in temperature, do not take space available, do not infinitely expand, medium to high density

Properties of liquids: not very compressible, flows – rate depends on viscosity, take shape of container, expand and contract with changes in temperature, not infinitely expandable, medium density

Properties of gases: very compressible, flow rapidly, take shape of container, expand and contract with changes in temperature, fill space available, infinitely expandable, exert greater pressure when temperature rises, low density

Gas variables: pressure, temperature, amount (moles, mass), temperature

Gas pressure

$$\underset{\text{Pa}}{\underset{\text{(Pascals)}}{\text{Pressure}}} = \frac{\text{Force}}{\text{area}}\ \frac{(\text{N})}{(\text{m}^2)}$$

Atmospheric pressure (arising from gravity attracting gases in the atmosphere to the Earth's surface) is measured with a barometer

1 atm	=	760 mm Hg
	=	760 Torr
	=	14.7 psi
	=	101,325 Pa
	=	1.01325 bar

Manometers measure pressure in the lab

Empirical gas laws:

1) **Boyle's law**: at constant T, the volume of a given mass is inversely proportional to the pressure of that gas → vary P (pressure) and V (volume) and keep T (temperature) and n (moles) constant

$$P_1V_1 = P_2V_2$$

2) **Charles' law:** at constant pressure, the volume of a given mass of has is directly proportional to the absolute temperature

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

3) **Combining Charles and Boyle's laws:** at constant volume, the pressure exerted by a given mass of gas is directly proportional to the absolute temperature

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

4) **Gay-Lussac and Avogadro's laws:**
   a. **Avogadro's law:** equal volumes of gases measured under the same conditions of temperature and pressure contain equal numbers of molecules (moles)

   $$\frac{V_1}{n_1} = \frac{V_2}{n_2}$$

   b. **Gay-Lussac's law:** law of combining volumes → when gases react, the ratio of the volumes and gases involved, measured at the same temperature and pressure, are small whole numbers.

These laws assume that intermolecular interactions are negligible.

The ideal gas law

$$PV = nRT$$

R is the universal gas constant and the value of R depends on the units of pressure

$$R = 0.082058 \text{ L atm K}^{-1}\text{ mol}^{-1}$$
$$= 8.3145 \text{ J K}^{-1}\text{ mol}^{-1}$$
$$= 8.3145 \text{ kg m}^{2}\text{ s}^{-1}\text{ K}^{-1}\text{ mol}^{-1}$$
$$= 8.3145 \text{ dm}^{3}\text{ kPa K}^{-1}\text{ mol}^{-1}$$

Standard temperature and pressure: 0 degrees (272.15 K) and 1 bar (1.00×10^5 Pa = 0.98 atm)

Standard molar volume: volume of 1 mole of gas at standard temperature and pressure – 22.7L

$$PV = \frac{\text{mass} \times RT}{M} \qquad \rho = \frac{P \times M}{RT}$$

Dalton's law of partial pressure: In a mixture of gases, the total pressure exerted is the sum of the partial pressures that each gas would exert if it alone were present under the same conditions

$$P_{tot} = P_A + P_B + P_C + \ldots\ldots$$

This is assuming no interactions → the gases are independent

Mole fraction: what percentage of the total number of gas molecules is a particular gas, the mole fraction = moles of a particular gas/ total number of moles

$X_A = n_A/n_{total}$

$P_A = X_A \times P_{tot}$

The pressure of each component ‘A’ of the mixture, is the mole fraction x total pressure

Molecular theory of gases

The gas laws describe the macroscopic behaviour of gases, however gases are made up of molecules

Kinetic molecular theory of gases

Gases are composed of molecules whose size is negligible compares with the average distance between them.

The molecules move randomly in straight lines, in all directions and at various speeds.

The forces of attraction or repulsion between two molecules in a gases are very weak or negligible, except when they collide.

The average kinetic energy of a molecule is proportional to the absolute temperature

Origin of pressure: force exerted when particles collide with a surface or eachother

Boyle’s law relating to pressure and volume: With less volume there is less space between the molecules and a shorter travel path → hence collisions occur more frequently → higher pressure

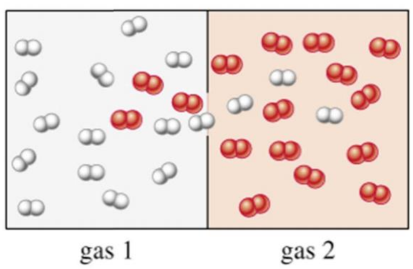
Charles’ law relating to temperature: with an increase in temperature, the kinetic and molecule speed increase – hence the collisions with the walls are more frequent and have more force → greater volume is occupied → higher pressure

Kinetic theory: different gases at the same temperature have the same average kinetic energy. Therefore this means that heavier gases travel more slowly with the same energy while lighter gases travel more quickly

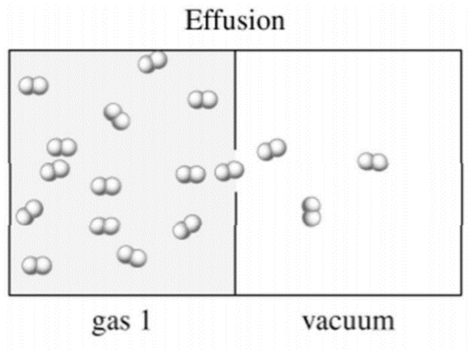
Diffusion and Effusion – Graham’s law

Diffusion: mixing of gases, until mixture is homogenous – the movement of particles from an area of high concentration to an area of low concentration until evenly distributed

Diffusion



Effusion: escape of molecules through a hole of molecular dimensions



Graham's law: the rate of effusion (or diffusion) of two gases at the same temperature ad pressure are inversely proportional to the square roots of their densities

$$\frac{\text{rate}_1}{\text{rate}_2} = \frac{\sqrt{M_2}}{\sqrt{M_1}} \qquad \frac{\text{time}_1}{\text{time}_2} = \frac{\sqrt{M_1}}{\sqrt{M_2}} \qquad \boxed{\frac{\text{rate}_1}{\text{rate}_2} = \frac{\sqrt{\rho_2}}{\sqrt{\rho_1}}}$$

Non-Ideal gases

Ideal gas: assumes zero particle volume and no interactions

No gas is deal - Volume is significant at higher pressures, interactions are more important at lower temperatures (attractive forces dominate)