GCE Physics Unit 3.3 Kinetic Theory

The mole:

The mole is the S.I. unit of an 'amount of substance'. It is the amount containing as many particles (e.g. molecules) as there are atoms in 12 g of carbon-12.

In 1 mole of water there are 6.02×10^{23} water molecules, this is the same number as of copper atoms in 1 mole of copper. It is known as the **Avogadro constant** and it represents **the number of particles in 1 mole of a substance**.

$$N_A = 6.02\times10^{23}\ \text{mol}^{-1}$$

Ideal gases:

Experiments investigating gas' behaviour discovered:

- Pressure $(p) \propto \frac{1}{\text{Volume }(V)}$ at a constant temperature (T)
- Pressure ∝ Temperature at a constant volume
- Volume ∝ Temperature at a constant pressure

These relationships can be combined into 1 equation:

$$pV=nRT$$

where n is the number of moles of gas and R is the molar gas constant = 8.31 J $\text{mol}^{-1}\ \text{K}^{-1}$.

Alternatively, it can be written as:

$$pV=NkT$$

where N is the number of particles and k is the Boltzmann constant = 1.38×10^{-23} J K^{-1}.

These equations are for ideal gases, but a very close approximations for the behaviour of real gases.

Kinetic theory:

In an ideal gas it is assumed that the molecules **collide with no loss of kinetic energy**, they only exert **forces on each other during collisions** and that the molecules are so small it can be assumed that they **take up no space**.

These assumptions also apply to the kinetic theory of gases:

- The time taken for a collision is negligible.
- The molecules move with constant velocity between collisions.
- The number of molecules is large, with a large number of collisions.
- The motion of the molecules is evenly distributed in all directions.
- **There is a random distribution of energy among the particles**.

Pressure:

The pressure a gas exerts is due to the motion of the molecules in the gas causing collisions with the surface of any container. Using the assumptions above the pressure can be calculated using these equations:

$$p = \frac{1}{3}\rho\overline{c^2} = \frac{1}{3}\frac{N}{V}m\overline{c^2}$$

where ρ is the density of the gas, N is the number of molecules and m **is the mass of one molecule**.

$\overline{c^2}$ is the mean squared speed of the molecules. This is important to use as the energy of the particles is distributed randomly, it is a way of representing the speed^2 of an average molecule.

Molar mass:

The relative molecular mass, M_r, of a molecule is its mass relative to the mass of carbon-12. The mass of 1 mole of a substance is known as its **molar mass**, M, it is linked to the M_r by this equation:

$$M/kg = \frac{M_r}{1000}$$

This can be used to calculate the **number of moles**, n, in a gas.

$$n = \frac{\text{mass of gas}}{\text{molar mass}}$$

Kinetic energy:

Combining these two equations can give an expression for the total translational kinetic energy of a gas.

$$pV = nRT \qquad p = \frac{1}{3}\frac{N}{V}m\overline{c^2}$$

$nRT = \frac{1}{3}Nm\overline{c^2}$ multiplying by $\frac{3}{2}$ gives

$\frac{3}{2}nRT = \frac{1}{2}Nm\overline{c^2}$ **as** $\boldsymbol{\frac{1}{2}m\overline{c^2}}$

is the mean translational kinetic energy for a molecule of a gas, multiplying this by the number of molecules will give the total translational kinetic energy for a gas. Therefore, **for 1 mole of a gas**,

Total translational kinetic energy $= \frac{3}{2}RT$

The mean kinetic energy of one molecule is

$$\frac{3}{2}kT$$

where $k = \frac{R}{N_A}$ = Boltzmann constant.

p = pressure in Pa | R = molar gas constant | N = number of molecules | k = Boltzmann constant | m = mass of one molecule in kg | M = molar mass in kg mol^{-1}

V = volume in m3 | T = temperature in K | n = number of moles | N_A = Avogadro constant | ρ = density in kg m^{-3} | M_r = relative molecular mass