

OCR B Physics A Level

Module 5.2: Matter Notes



5.2.1 Matter: Very Simple

Gas Behaviour

Gas particles are in constant, random, **Brownian** motion. They continuously collide with each other and the walls of their container, causing a **change in momentum** (or **impulse**), which produces a force. This force results in **pressure** ($P = F/A$). This is how gas particles exert pressure on their containers.

Impulse is equal to the **area under a force-time graph**.

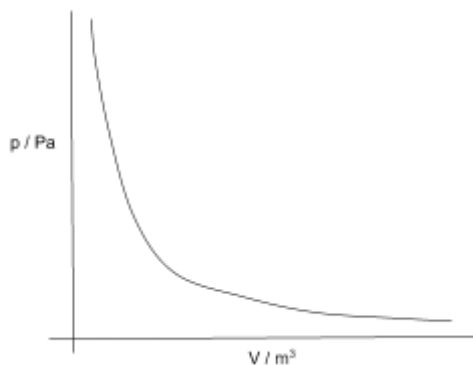
Particles take a **random path** - they do not travel in a straight line, but are constantly changing direction due to collisions. This is why processes like diffusion are so slow.

$$\text{displacement} \propto \sqrt{\text{number of steps}}$$

The Gas Laws

Boyle's Law: at a constant temperature, pressure p and volume V are inversely proportional:

$pV = \text{constant}$. This is because, as volume decreases, the particles become closer together (more densely packed) so collide more frequently, exerting a greater force and therefore pressure on the container.



Charles' Law: at a constant pressure, volume V is directly proportional to absolute temperature T .

The Pressure Law: at a constant volume, pressure p is directly proportional to absolute temperature T .

Ideal Gases

In Kinetic Theory, gases are assumed to be **ideal gases**. This is based on the following assumptions:

- The gas contains a large number of molecules
- The molecules are identical to each other
- All collisions between molecules and the walls of their container are **perfectly elastic** (in other words, energy is conserved)
- The time taken for collisions is negligible compared to the time between collisions
- There are **no intermolecular forces**, so molecules do not attract each other
- Molecules are in constant, random (**Brownian**) motion
- The gas particles obey Newton's Laws of motion



The Ideal Gas Equations

$$pV = nRT$$

Where...

- p = pressure (Pa)
- V = volume (m^3)
- n = number of moles (mol)
- R = gas constant, $8.314 \text{ J mol}^{-1} \text{ K}^{-1}$
- T = temperature (K)

$$pV = NkT$$

Where...

- p = pressure (Pa)
- V = volume (m^3)
- N = number of molecules
- k = Boltzmann constant, $1.38 \times 10^{-23} \text{ (J K}^{-1}\text{)}$
- T = temperature (K)

The ideal gas laws are derived by combining all three gas laws:

$$pV = \frac{1}{3} Nm \overline{c^2}$$

Consider a particle, mass m , in a cubic box, with sides of length x .

The particle can move in **3 dimensions**, so the pressure exerted in any one dimension is $\frac{1}{3}$ of the total pressure, hence the fraction in the equation.

A particle colliding with the wall will transfer momentum $2mc$ (where c is the speed of the particle), because its velocity after is equal and opposite to its velocity before (as energy is conserved in elastic collisions).

Force is equal to change in momentum/time, and the time between two consecutive collisions on the same wall is equal to: $\frac{2x}{c}$

Therefore, force is equal to:

$$\frac{2mc}{2x/c} = \frac{mc^2}{x}$$

Pressure is force divided by area, so pressure is equal to

$$\frac{mc^2}{x} \div x^2 = \frac{mc^2}{x^3}$$

x^3 is equal to volume, so this can be rewritten as:

$$p = \frac{mc^2}{V}$$

and rearranged to give:

$$pV = mc^2$$

This equation is for a single particle; multiply by $\frac{1}{3} N$ to find the pressure exerted by all of the particles on one wall. This gives the final equation:

$$pV = \frac{1}{3} Nmc^2$$



Energy

These equations can be related to the energy of a system.

The **internal energy** of a system is the sum of the kinetic and potential energies of its particles. However, for an ideal gas, there are assumed to be no intermolecular forces, and therefore no potential energy, so the assumption that internal energy = kinetic energy is made.

Rearranging the ideal gas equation for this purpose gives:

$$NkT = \frac{1}{3}Nmc^2 \rightarrow kT = \frac{1}{3}mc^2 \rightarrow \frac{3}{2}kT = \frac{1}{2}mc^2 \text{ (where } \frac{1}{2}mc^2 \text{ is kinetic energy)}$$

Therefore kinetic energy is approximated to $\frac{3}{2}kT$ or $\frac{3RT}{2N_A}$.

Energy Changes

Specific thermal capacity or **specific heat capacity** is the amount of energy needed to raise the temperature of 1kg of a material/substance by 1°C.

This can be used to work out the energy change when heating an object:

$$E = mc\Delta\theta$$

Where E = energy / J

m = mass / kg

θ = temperature / t

You can determine the specific heat capacity of a material by heating it, and measuring the energy input alongside its temperature change. Plot a graph of temperature change against energy, calculate the gradient and divide by mass. This will give you 1/c.



5.2.2 Matter Hot and Cold

Temperature and Energy

Remember: $E = \frac{3}{2} kT$

Chemical processes require a certain amount of energy in order to take place. This is known as the **activation energy**. If $\frac{3}{2} kT$ is less than the activation energy, the reaction will be unable to occur.

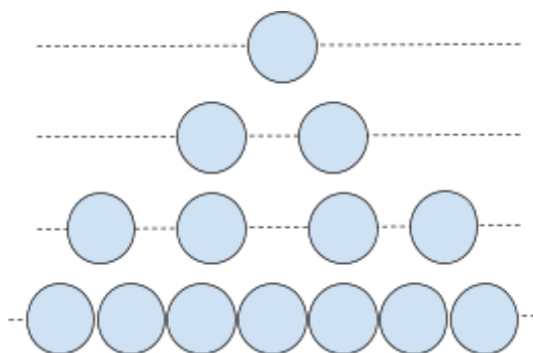
When the activation energy is between $15kT$ and $30kT$, reactions occur at an **appreciable rate**.

Probability and the Boltzmann Factor

Particles exchange energy through **random collisions**. The probability of a particle gaining enough energy to move to a higher energy level is determined by these random collisions.

The number of particles on each level decreases exponentially with each level up.

The **Boltzmann factor** is the ratio of the number of particles on a lower level to the number on a higher level.



Eg. The Boltzmann factor here is $\frac{1}{2}$ because this is the ratio of particles between adjacent levels.

The Boltzmann factor depends on the process involved; it can be the movement of a particle from one level to the level above, or to two levels above, and the ratio is based on the “jump” involved.

The Boltzmann factor is also given by the following equation.

$$Bf = e^{-\frac{\epsilon}{kT}}$$

Where Bf = Boltzmann factor

e = exponential factor

ϵ = activation energy

k = Boltzmann constant, 1.4×10^{-23}

T = temperature

