AQA Physics A-level

Section 6.2: Thermal Physics

Notes

3.6.2 Thermal physics

3.6.2.1 - Thermal energy transfer

The **internal energy** of a body is equal to the sum of all of the kinetic energies and potential energies of **all** its particles. The kinetic and potential energies of a body are **randomly distributed**.

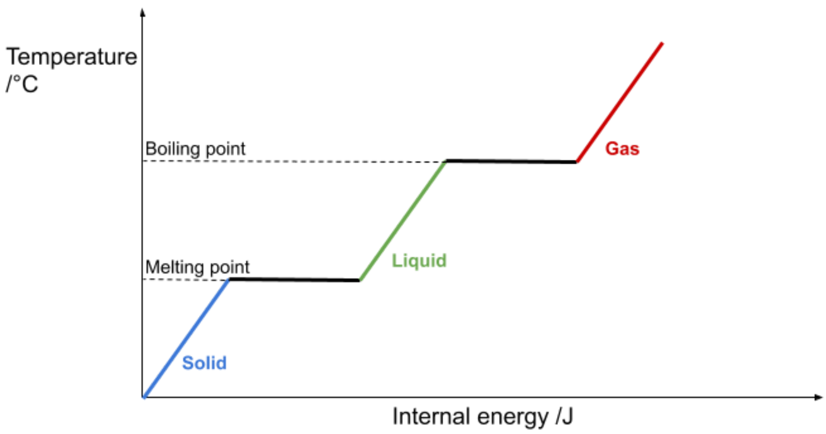
The **internal energy** of a system can be **increased** in two ways:

- **Do work** on the system to transfer energy to it, (e.g moving its particles/changing its shape).
- **Increase the temperature** of the system.

When the state of a substance is changed, its internal energy also changes, this is because the **potential energy of the system changes, while the kinetic energy of the system is kept constant**. This can be demonstrated by measuring the temperature of water as it boils:

- The temperature increases up until 100°C, after which the energy gained through heating the water is no longer used to increase the temperature (and therefore kinetic energy), but instead is used to break bonds between water molecules so it can change state to water vapour, and so the potential energy is increased.

Below is a graph showing how the internal energy of a substance varies with temperature:



You can measure the amount of energy required to change the temperature of a substance using the following formula: $Q = mc\Delta\theta$ Where Q is energy required, m is the mass, c is the specific heat capacity, and $\Delta\theta$ is the change in temperature.

The **specific heat capacity** of a substance is the amount of energy required to increase the temperature of 1 kg of a substance by 1 °C/1 K, without changing its state.

You can measure the amount of energy required to change the state of a substance using the following formula: $Q = ml$

Where Q is energy required and l is the specific latent heat.

The **specific latent heat** of a substance is the amount of energy required to change the state of 1 kg of material, without changing its temperature. There are two types of specific latent heat: the **specific latent heat of fusion** (when solid changes to liquid) and **specific latent heat of vaporisation** (when liquid changes to gas).

You will need to be able to do calculations using the above formulas and understand **continuous-flow** questions, below are some examples:

A kettle has a power of 1200 W, and contains 0.5 kg of water at 22°C, how long will it take for the water in the kettle to reach 100°C? (specific heat capacity of water = 4200 J/kg°C)

Firstly, you must find the energy required to increase the temperature of the water to 100°C using $Q = mc\Delta\theta$.

$Q = 0.5 \times 4200 \times (100 - 22) = 2100 \times 78 =$ **163800 J**

Power is the energy transferred over time, therefore to find the value of time taken we must divide the energy required by the power.

$P = \frac{Q}{t} \rightarrow t = \frac{Q}{P}$ $\quad t = \frac{163800}{1200} =$ **136.5 s**

An ice cube of mass 0.01 kg at a temperature of 0°C is dropped into a glass of water of mass 0.2 kg, at a temperature of 19°C. What is the final temperature of the water once the ice cube has fully melted? (specific heat capacity of water = 4200 J/kg°C, specific latent heat of fusion of ice = 334 $\times 10^3$ J/kg)

Firstly, find the energy required to change the state of the ice.

$Q = ml$ $\quad Q = 0.01 \times 334 \times 10^3$ **= 3340 J**

Next, you must set up a pair of simultaneous equations to show the energy transfer in the water and in the ice separately. As we know that the energy transfer is the same in both as the system is closed, we can equate these values to find the final temperature (***T***).

For ice: $Q = ml + mc\Delta\theta$ (because the ice changes state and temperature)

$Q = 3340 + 0.01 \times 4200 \times (T - 0)$ $\quad Q = 3340 + 42T$

For water: $Q = mc\Delta\theta$

$Q = 0.2 \times 4200 \times (19 - T)$ $\quad Q = 15960 - 840T$

Set them equal: $3340 + 42T = 15960 - 840T$

$882T = 12620$

$T =$ **14.3 °C**

Water flows past an electric heater with a power of 9000W at a rate of 0.5 kg/s. What is the increase in temperature of the water per second that it flows past the heater? (specific heat capacity of water = 4200 J/kg°C)

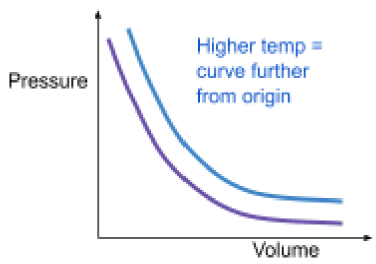
The power of the heater is 9000W, therefore we know 9000 J of energy are transferred every second, we also know 0.5 kg of water flows past the heater in that second, using these pieces of information we can find the increase in temperature:

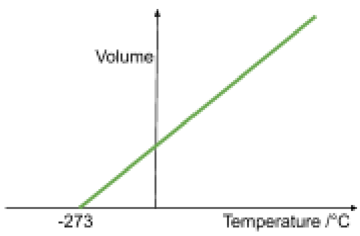
$Q = mc\Delta\theta$ $\quad \Delta\theta = \frac{Q}{mc}$ $\quad \Delta\theta = \frac{9000}{0.5 \times 4200} =$ **4.3°C**

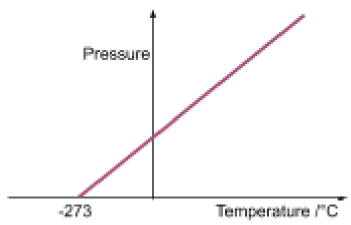
3.6.2.2 - Ideal gases

The **gas laws** describe the **experimental relationship between pressure (p), volume (V), and temperature (T) for a fixed mass of gas**. They are **empirical** in nature, meaning they are not based on theory but arose from observation and experimental evidence. The 3 gas laws you need to be aware of are:

1. **Boyle's Law** -When temperature is constant, pressure and volume are inversely proportional $$pV = k$$
2. **Charles' Law** -When pressure is constant, volume is directly proportional to absolute temperature $$\frac{V}{T} = k$$
3. **The Pressure Law** -When volume is constant, pressure is directly proportional to absolute temperature $$\frac{p}{T} = k$$







The absolute scale of temperature is the **kelvin scale**. All equations in thermal physics will use temperature measured in kelvin (K). A change of 1 K is equal to a change of 1°C , and to convert between the two you can use the formula:

$$K = C + 273$$

Where K is the temperature in kelvin and C is the temperature in Celsius.

Absolute zero (- 273°C), also known as 0 K, is the lowest possible temperature, and is the temperature at which particles have **no kinetic energy** and the volume and pressure of a gas are zero.

You can combine all the experimental gas laws into one to get $\frac{pV}{T} = k$ where the constant k is dependent on the amount of gas used measured in **moles**, therefore you can rewrite the above equation to get $\frac{pV}{T} = nR$, where ***n*** is the number of moles of gas, and ***R*** is the molar gas constant (8.31 J mol^{-1} K^{-1}). You can rearrange this further to get:

$$pV = nRT$$, which is the ideal gas equation.

1 mole of a substance is equal to 6.02×10^{23} atoms/molecules, so you can convert between the number of moles (n) and the number of molecules (N) by multiplying the number of moles by 6.02×10^{23}, which is defined as the Avogadro constant (N_A).

$$N = n \times N_A \quad \Rightarrow n = \frac{N}{N_A}$$

You can substitute in the above equation into the ideal gas equation to get it in terms of molecules rather than moles: $pV = \frac{NRT}{N_A}$. You can simplify this further by using the Boltzmann constant (k), which is equivalent to $\frac{R}{N_A}$, leading to the equation:

$$pV = NkT$$

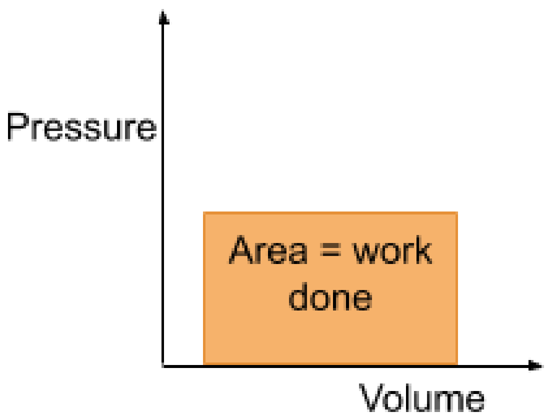
Molar mass is the mass (in grams) of one mole of a substance and can be found by finding the **relative molecular mass**, which is (approximately) equal to the sum of the nucleons in a molecule of the substance.

Work is done on a gas to change its volume when it is at constant pressure, (this is usually done through the transfer of thermal energy) the value of work done can be calculated using the formula:

$$Work\ done = p\Delta V$$

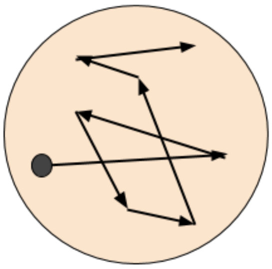
Where p is the pressure and ΔV is the change in volume.

This is especially useful when working with a graph of pressure against volume, as work done is simply the area under the graph.



3.6.2.3 - Molecular kinetic theory model

Brownian motion is the random motion of larger particles in a fluid caused by **collisions** with surrounding particles, and can be observed through looking at smoke particles under a microscope. Brownian motion contributed to the evidence for the existence of atoms and molecules.



You can use a **simple molecular model** to explain each of the gas laws:

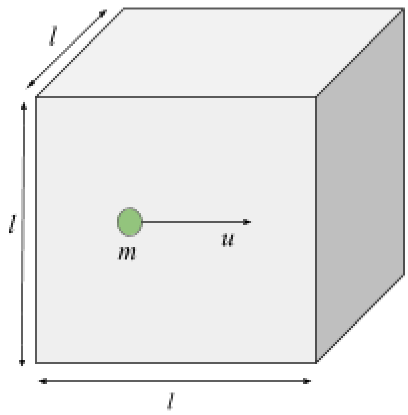
- **Boyle's law** - Pressure is inversely proportional to volume at constant temperature
  E.g If you increase the volume of a fixed mass of gas, its molecules will move further apart so collisions will be less frequent therefore pressure decreases.
- **Charles's law** - Volume is directly proportional to temperature at constant pressure
  When the temperature of a gas is increased, its molecules gain kinetic energy meaning they will move more quickly and because pressure is kept constant (therefore frequency of collisions is constant) the molecules move further apart and volume is increased.
- **Pressure Law** - Pressure is directly proportional to temperature at constant volume
  When the temperature of a gas is increased, its molecules gain kinetic energy meaning they will move more quickly, as volume is constant the frequency of collisions between molecules and their container increases and they collide at higher speeds therefore pressure is increased.

The above laws are **empirical** in nature, meaning they are not based on theory but arose from observation and experimental evidence. The **kinetic theory model** however is the opposite and arose from only theory.

The **kinetic theory model** equation relates several features of a fixed mass of gas, including its pressure, volume and mean kinetic energy. There are several underlying **assumptions**, which lead to the derivation of this equation, these assumptions and the derivation are outlined below.

Assumptions -

- **No intermolecular forces** act on the molecules
- The **duration of collisions is negligible** in comparison to time between collisions
- The motion of molecules in **random**, and they experience **perfectly elastic collisions**
- The motion of the molecules follows **Newton's laws**
- The molecules **move in straight lines** between collisions



Derivation -

1. First, you must consider a cube with side lengths l, full of gas molecules. One of these molecules, has a mass m and is travelling towards the right-most wall of the container, with a velocity u. Assuming it collides with this wall elastically, its **change in momentum** is $mu - (-mu) = 2mu$.

2. Before this molecule can collide with this wall again it must travel a distance of $2l$. Therefore the time between collisions is t, where $t = \frac{2l}{u}$.

3. Using these two bits of information we can find the **impulse**, which is the rate of change of momentum of the molecule. As impulse is equal to the **force** exerted, we can find **pressure** by dividing our value of impulse by the area of one wall: l^2.

$$F = \frac{2mu}{\frac{2l}{u}} = \frac{mu^2}{l} \qquad P = \frac{mu^2}{\frac{l}{l^2}} = \frac{mu^2}{l^3} = \frac{mu^2}{V}$$

As shown, the above equation can be further simplified because l^3 is equal to the cube's volume (V).

4. The molecule we have considered is one of many in the cube, the total pressure of the gas will be the sum of all the individual pressures caused by each molecule.

$$P = \frac{m((u_1)^2+(u_2)^2+...+(u_n)^2)}{V}$$

5. Instead of considering all these speeds separately, we can define a quantity known as **mean square speed**, which is exactly what it sounds like, the mean of the square speeds of the gas molecules. This quantity is known as $\overline{u^2}$, and we multiply it by N, the number of particles in the gas, to get an estimate of the sum of the molecules' speeds.

$$P = \frac{Nm\overline{u^2}}{V}$$

6. The last step is to consider all the directions the molecules will be moving in. Currently we have only considered one dimension, however the particles will be moving in all 3 dimensions. Using pythagoras' theorem we can work out the speed the molecules will be travelling at:

$$\overline{c^2} = \overline{u^2} + \overline{v^2} + \overline{w^2}$$

Where u, v, and w are the components of the molecule's velocity in the x, y and z directions.
As the motion of the particles is random we can assume the mean square speed in each direction is the same.

$$\overline{u^2} = \overline{v^2} = \overline{w^2} \quad \therefore \quad \overline{c^2} = 3\overline{u^2}$$

The last thing to do now it put this into our equation and rearrange:

$$pV = \frac{1}{3}Nm\overline{c^2} \quad \text{or} \quad pV = \frac{1}{3}Nm(c_{rms})^2$$

As $\overline{c^2}$ and $(c_{rms})^2$ are equivalent

An **ideal gas** follows the gas laws perfectly, meaning that there is no other interaction other than **perfectly elastic collisions** between the gas molecules, which shows that no intermolecular forces act between molecules. As potential energy is associated with intermolecular forces, **an ideal gas has no potential energy**, therefore its internal energy is equal to the sum of the kinetic energies of all of its particles.

There are several equations which allow you to find the kinetic energy of a **single gas molecule**, therefore these can be used to find the internal energy of an ideal gas:

$$\frac{1}{2}m(c_{rms})^2 = \frac{3}{2}kT = \frac{3RT}{2N_A}$$

As you can see from the middle equation, the kinetic energy of a gas molecule is directly proportional to temperature (in Kelvin).

Below is an example question using the above equations, as well as knowledge of molar mass and molecular mass from the previous section:

A bottle contains 128 g of oxygen at a temperature of 330 K. Find the sum of the kinetic energies of all the oxygen molecules. (Molecular mass of oxygen gas = 32 g)

Firstly, find the number of moles of gas, then multiply this by the avogadro constant to find the number of molecules.

Number of moles = $\frac{mass}{molar\ mass} = \frac{128}{32} = 4$

Number of molecules = $4 \times 6.02 \times 10^{23} = 2.408 \times 10^{24}$

Then, use $\frac{3}{2}kT$ to find the kinetic energy of one molecule and then multiply this by the number of molecules:

Kinetic energy of a single molecule = $\frac{3}{2} \times 1.38 \times 10^{-23} \times 330$ = 6.831×10^{-21}

Sum of kinetic energies = $6.831 \times 10^{-21} \times 2.408 \times 10^{24}$ = 16450 J

Knowledge and understanding of gases has changed greatly over time; the gas laws were discovered by a number of scientists and later explained by the development of the kinetic theory model, however this model wasn't accepted at first. Knowledge and understanding of any scientific concept changes over time in accordance to the experimental evidence gathered by the scientific community.