

Change of temperature and state

A **change of temperature** will occur when a substance is heated or cooled, and, as a result, the **kinetic energy** of the molecules will change and, therefore, so will the **internal energy**.

The **heat inputted or removed** from a system, Q , is

$$Q = mc\Delta\theta$$

m = mass, $\Delta\theta$ = change in temperature, c = specific heat capacity

A substance will **change state** when provided with enough energy.

During a change of state the **temperature doesn't change** and, therefore, the **kinetic energy** remains constant. However, the **potential energy** of the molecules changes and, therefore, so does the **internal energy**.

The **change in heat required for a change of state** is:

$$Q = ml$$

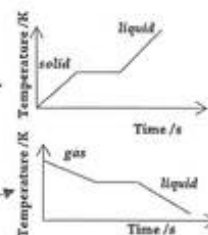
l = specific latent heat

l_f is the **specific latent heat of fusion**, the specific latent heat for a change from a solid to a liquid, or vice versa

l_v is the **specific latent heat of vaporisation**, the specific latent heat for a change from a liquid to a gas, or vice versa

Solid to liquid

Gas to liquid



Thermal energy transfer

Internal energy: the sum of the randomly distributed kinetic energies and potential energies of particles in a system or body.

First law of thermodynamics

The change of internal energy of the object is equal to the total energy transfer due to work done and heating.

$$\Delta U = \Delta Q + \Delta W$$

ΔU = change in internal energy, ΔQ = change in heat, ΔW = work done

Internal energy is constant when

- no energy transfer by heating and no work is done OR
- energy transfer by heating and work done are equal and opposite

Internal energy can be increased by:

- heating the object OR
- doing work on the object

Imagine an engine – as the air is compressed, it heats up. No heat is being added, but work is being done – this raises the internal energy, so temperature increases.

Exam Tip!

Keep track of direction of energy transfer. Heat flowing **into** an object means ΔQ is **positive**, while heat flowing **out** of an object means ΔQ is **negative**.

Thermal physics

Gas laws: a set of experimental laws that look at the relationship between pressure, volume and temperature of a gas.
Molar mass: the mass of one mole of substance in kg mol^{-1}
Molecular mass: the mass of a molecule in u

Ideal gas law:

$$pV = nRT$$

The pressure law (constant volume):

$$\frac{p}{T} = \text{constant}$$

Charles' law (constant pressure):

$$\frac{V}{T} = \text{constant}$$

Boyle's law (constant temperature):

$$pV = \text{constant}$$

Work done on a gas:

$$W = p\Delta V$$

Avogadro's constant (N_A) is the number of atoms in 12 g of carbon isotope $^{12}_6\text{C}$.
An isotope's or chemical's **molecular mass** describes how much mass is in N_A molecules of that chemical.

The **Boltzmann constant** (k) is related to

Avogadro's constant by $k = \frac{R}{N_A}$

p = pressure of the gas

V = volume of the gas

T = temperature of the gas

n = number of moles in a gas

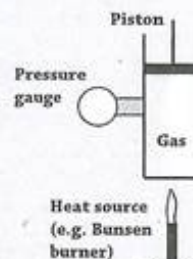
W = work done on the gas

N_A = Avogadro's constant = $6.02 \times 10^{23} \text{ mol}^{-1}$

R = gas constant = $8.31 \text{ J K}^{-1} \text{ mol}^{-1}$

k = Boltzmann constant = $1.38 \times 10^{-23} \text{ J K}^{-1}$

Ideal gas laws



This set-up can be used to investigate the gas laws.

A **Bunsen burner** allows the temperature to be controlled and a **piston** allows the volume to be controlled.

By measuring the pressure and controlling both temperature and volume, pressure can also be kept constant.



Think of a balloon. By compressing it, the **volume decreases** and the **pressure increases** while the **temperature stays the same**.

If we increased the temperature while holding the volume constant, the pressure would increase.

Molecular kinetic theory model

The molecular kinetic theory applies the gas laws to **individual molecules**. Molecules in a fluid are in **constant random motion**. The kinetic theory is based on mathematics.

Brownian motion describes the erratic motion of a small particle when placed on the surface of a fluid. The molecules on the surface of the fluid move randomly on the surface and hit into the particle, making the particle move randomly on the surface – this produces Brownian motion, providing evidence for the molecular kinetic theory.

Molecular kinetic theory model representations of the gas laws

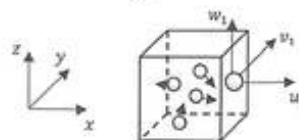
- Pressure law:** As T is increased in a container with fixed V , the **average speed of molecules increases**. This causes molecules to impact the walls of a container with **greater force**, and p increases. $p \propto T$
- Charles' law:** As T is increased in a flexible container, the **faster moving molecules** impact the walls of the container with **greater force**. If the walls are movable, this increases V while p remains the same. $V \propto T$
- Boyle's law:** In a case where T is constant, but V is decreased, molecules have less space to move in between collisions. This **increases the number of collisions**, so the total force on the walls of the container increases, increasing p . $p \propto 1/V$

Kinetic theory equation

The kinetic theory equation is derived by considering one molecule of mass m in a rectangular box of dimensions l_x, l_y, l_z with speeds u, v, w in x, y, z directions.

1. The force of an impact of a particle on a wall is the change in momentum as it collides

$$F_1 = +\frac{mu_1^2}{l_x}$$



2. $p = F/A$ so

$$p_1 = \frac{mu_1^2}{l_x l_y l_z} = \frac{mu_1^2}{V}$$

3. The total pressure on the box is the sum of pressure of each individual particle

$$p = p_1 + p_2 + p_3 + \dots + p_N$$

$$p = \frac{m}{V} (u_1^2 + u_2^2 + u_3^2 + \dots + u_N^2)$$

The average squared velocity ($\overline{u^2}$) is given by $\frac{u_1^2 + u_2^2 + u_3^2 + \dots + u_N^2}{N}$

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

4. This is only for one dimension so add the other two

$$p = \frac{Nm}{3V} (\overline{u^2} + \overline{v^2} + \overline{w^2})$$

Collecting dimensions together gives:

$$p = \frac{Nm}{3V} c_{rms}^2$$

c_{rms}^2 = root mean square velocity

Average molecular kinetic energy:

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

Internal energy of an ideal gas:

For an ideal gas the internal energy is equal to the kinetic energy, as potential energy = 0

$$N_A \times n \times \frac{3}{2} kT = \frac{3}{2} nRT$$

Key assumptions

- Gas molecules move randomly
- Volume of gas molecules is negligible
- There are no intermolecular forces between molecules
- Time spent between collisions is significantly longer than duration of collisions
- All collisions are elastic