

3 THERMAL PHYSICS

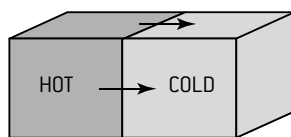
Thermal concepts

TEMPERATURE AND HEAT FLOW

Hot and cold are just labels that identify the direction in which thermal energy (sometimes known as heat) will be naturally transferred when two objects are placed in thermal contact. This leads to the concept of the 'hotness' of an object. The direction of the natural flow of thermal energy between two objects is determined by the 'hotness' of each object. Thermal energy naturally flows from hot to cold.

The temperature of an object is a measure of how hot it is. In other words, if two objects are placed in thermal contact, then the temperature difference between the two objects will determine the direction of the natural transfer of thermal energy. Thermal energy is naturally transferred 'down' the temperature difference – from high temperature to low temperature. Eventually, the two objects would be expected to reach the same temperature. When this happens, they are said to be in **thermal equilibrium**.

Heat is not a substance that flows from one object to another. What has happened is that thermal energy has been transferred. Thermal energy (heat) refers to the non-mechanical transfer of energy between a system and its surroundings.



direction of transfer of thermal energy

KELVIN AND CELSIUS

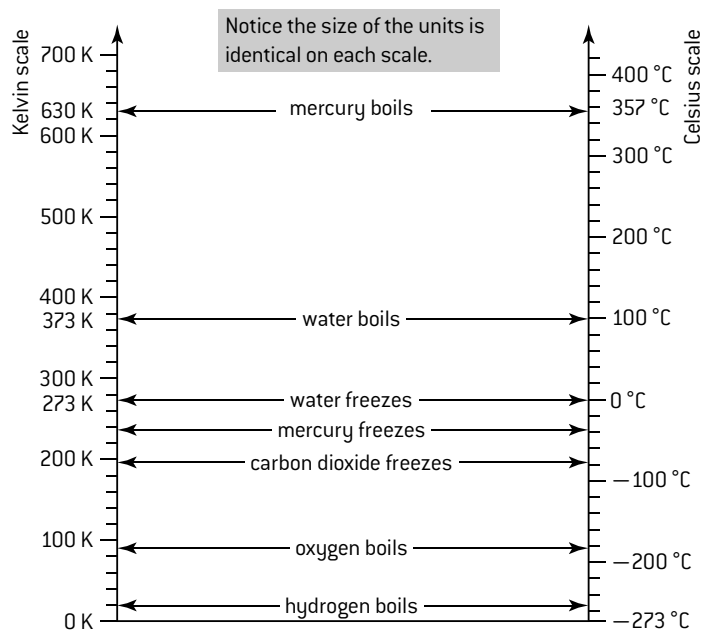
Most of the time, there are only two sensible temperature scales to choose between – the Kelvin scale and the Celsius scale.

In order to use them, you do not need to understand the details of how either of these scales has been defined, but you do need to know the relation between them. Most everyday thermometers are marked with the Celsius scale and temperature is quoted in degrees Celsius ($^{\circ}\text{C}$).

There is an easy relationship between a temperature T as measured on the Kelvin scale and the corresponding temperature t as measured on the Celsius scale. The approximate relationship is

$$T (\text{K}) = t (^{\circ}\text{C}) + 273$$

This means that the 'size' of the units used on each scale is identical, but they have different zero points.



The Kelvin scale is an absolute thermodynamic temperature scale and a measurement on this scale is also called the *absolute temperature*.

Zero Kelvin is called *absolute zero* (see page 29).

EXAMPLES: GASES

For a given sample of a gas, the *pressure*, the *volume* and the *temperature* are all related to one another.

- The pressure, P , is the force per unit area from the gas acting at 90° on the container wall.

$$p = \frac{F}{A}$$

The SI units of pressure are N m^{-2} or Pa (Pascals).

$$1 \text{ Pa} = 1 \text{ N m}^{-2}$$

Gas pressure can also be measured in atmospheres

$$(1 \text{ atm} \approx 10^5 \text{ Pa})$$

- The volume, V , of the gas is measured in m^3 or cm^3 ($1 \text{ m}^3 = 10^6 \text{ cm}^3$)
- The temperature, t , of the gas is measured in $^{\circ}\text{C}$ or K

In order to investigate how these quantities are interrelated, we choose:

- one quantity to be the independent variable (the thing we alter and measure)
- another quantity to be the dependent variable (the second thing we measure).
- The third quantity needs to be controlled (i.e. kept constant). The specific values that will be recorded also depend on the mass of gas being investigated and the type of gas being used so these need to be controlled as well.

Heat and internal energy

MICROSCOPIC VS MACROSCOPIC

When analysing something physical, we have a choice.

- The **macroscopic** point of view considers the system as a whole and sees how it interacts with its surroundings.
- The **microscopic** point of view looks inside the system to see how its component parts interact with each other.

So far we have looked at the temperature of a system in a macroscopic way, but all objects are made up of **atoms and molecules**.


According to **kinetic theory** these particles are constantly in random motion – hence the name. See below for more details. Although atoms and molecules are different things (a molecule is a combination of atoms), the difference is not important at this stage. The particles can be thought of as little ‘points’ of mass with velocities that are continually changing.

INTERNAL ENERGY

If the temperature of an object changes then it must have gained (or lost) energy. From the microscopic point of view, the molecules must have gained (or lost) this energy.

The two possible forms are kinetic energy and potential energy.

speed in a random direction $\rightarrow v$
 \therefore molecule has KE

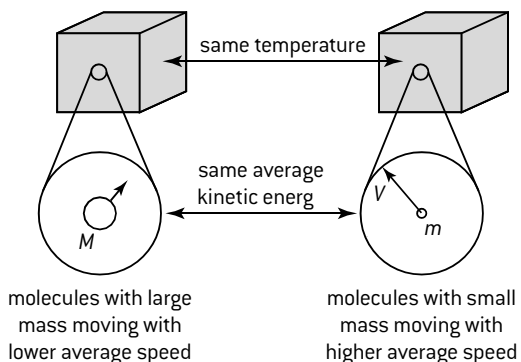
equilibrium position \ast $F \leftarrow$ 
 resultant force back towards equilibrium position due to neighbouring molecules
 \therefore molecule has PE

- The molecules have kinetic energy because they are moving. To be absolutely precise, a molecule can have either translational kinetic energy (the whole molecule is moving in a certain direction) or rotational kinetic energy (the molecule is rotating about one or more axes).
- The molecules have potential energy because of the **intermolecular** forces. If we imagine pulling two molecules further apart, this would require work against the intermolecular forces.

The total energy that the molecules possess (random kinetic plus inter-molecular potential) is called the **internal energy** of a substance. Whenever we heat a substance, we increase its internal energy.

Temperature is a measure of the average kinetic energy of the molecules in a substance.

If two substances have the same temperature, then their molecules have the same average kinetic energy.

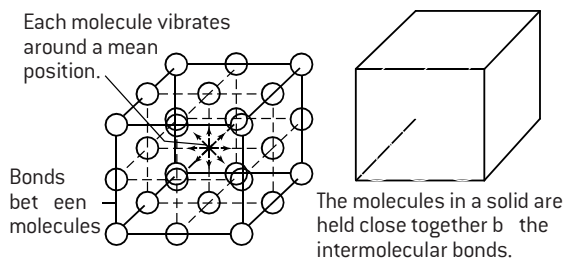


KINETIC THEORY

Molecules are arranged in different ways depending on the **phase** of the substance (i.e. solid, liquid or gas).

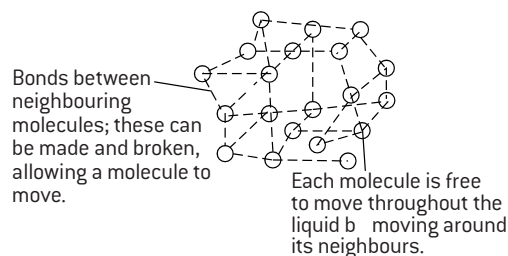
SOLIDS

Macroscopically, solids have a fixed volume and a fixed shape. This is because the molecules are held in position by bonds. However the bonds are not absolutely rigid. The molecules vibrate around a mean (average) position. The higher the temperature, the greater the vibrations.



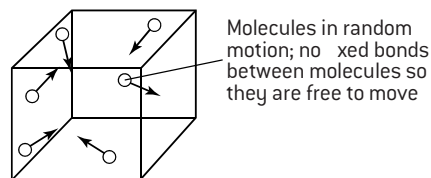
LIQUIDS

A liquid also has a fixed volume but its shape can change. The molecules are also vibrating, but they are not completely fixed in position. There are still strong forces between the molecules. This keeps the molecules close to one another, but they are free to move around each other.



GASES

A gas will always expand to fill the container in which it is put. The molecules are not fixed in position, and any forces between the molecules are very weak. This means that the molecules are essentially independent of one another, but they do occasionally collide. More detail is given on page 31.



HEAT AND WORK

Many people have confused ideas about heat and work. In answers to examination questions it is very common to read, for example, that ‘heat rises’ – when what is meant is that the transfer of thermal energy is upwards.

- When a force moves through a distance, we say that work is done. Work is the energy that has been transmitted from one system to another from the macroscopic point of view.
- When work is done on a microscopic level (i.e. on individual molecules), we say that heating has taken place. Heat is the energy that has been transmitted. It can either increase the kinetic energy of the molecules or their potential energy or, of course, both.

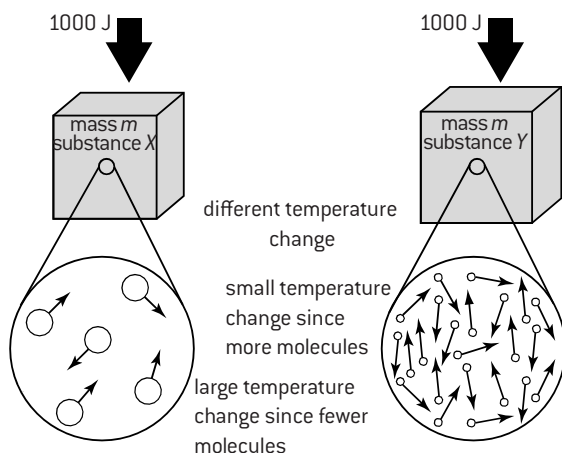
In both cases energy is being transferred.

Specific heat capacity

DEFINITIONS AND MICROSCOPIC EXPLANATION

In theory, if an object could be heated up with no energy loss, then the increase in temperature ΔT depends on three things:

- the energy given to the object Q ,
- the mass, m , and
- the substance from which the object is made.



Two different blocks with the same mass and same energy input will have a different temperature change.

We define the **thermal capacity** C of an object as the energy required to raise its temperature by 1 K. Different objects (even different samples of the same substance) will have different values of heat capacity. **Specific heat capacity** is the energy required to raise a unit mass of a substance by 1 K. 'Specific' here just means 'per unit mass'.

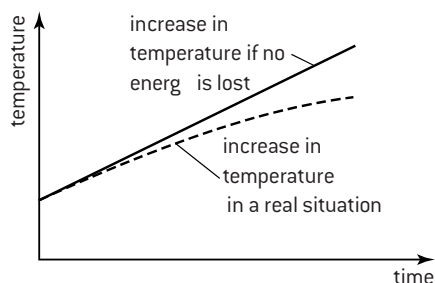
In symbols,

Thermal capacity $C = \frac{Q}{\Delta T}$ (J K^{-1} or $\text{J }^\circ\text{C}^{-1}$)

Specific heat capacity $c = \frac{Q}{(m \Delta T)}$ ($\text{J kg}^{-1} \text{K}^{-1}$ or $\text{J kg}^{-1} \text{ }^\circ\text{C}^{-1}$)
 $Q = mc\Delta T$

Note

- A particular gas can have many different values of specific heat capacity – it depends on the conditions used – see page 161.
- These equations refer to the **temperature difference** resulting from the addition of a certain amount of energy. In other words, it generally takes the same amount of energy to raise the temperature of an object from 25 °C to 35 °C as it does for the same object to go from 402 °C to 412 °C. This is only true so long as energy is not lost from the object.
- If an object is raised above room temperature, it starts to lose energy. The hotter it becomes, the greater the rate at which it loses energy.



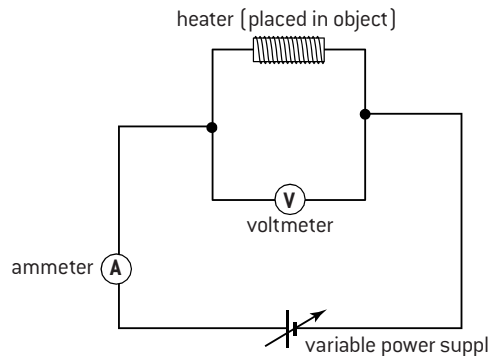
Temperature change of an object being heated at a constant rate

METHODS OF MEASURING HEAT CAPACITIES AND SPECIFIC HEAT CAPACITIES

There are two basic ways to measure heat capacity.

1. Electrical method

The experiment would be set up as below:



• the specific heat capacity $c = \frac{I t V}{m(T_2 - T_1)}$

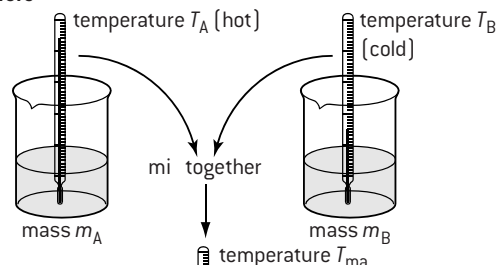
Sources of experimental error

- loss of thermal energy from the apparatus.
- the container or the substance and the heater will also be warmed up.
- it will take some time for the energy to be shared uniformly through the substance.

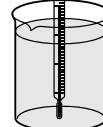
2. Method of mixtures

The known specific heat capacity of one substance can be used to find the specific heat capacity of another substance.

before



after



Procedure:

- measure the masses of the liquids m_A and m_B .
- measure the two starting temperatures T_A and T_B .
- mix the two liquids together.
- record the maximum temperature of the mixture T_{max} .

If no energy is lost from the system then,

energy lost by hot substance cooling down = energy gained by cold substance heating up

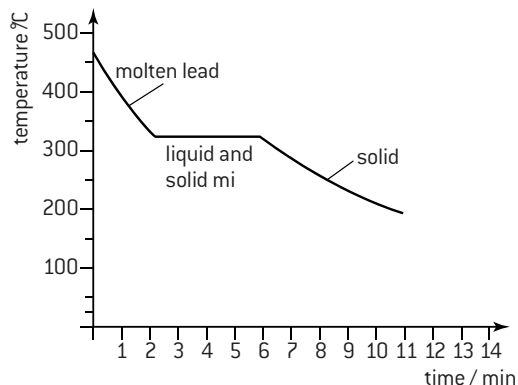
$$m_A c_A (T_A - T_{\text{max}}) = m_B c_B (T_{\text{max}} - T_B)$$

Again, the main source of experimental error is the loss of thermal energy from the apparatus; particularly while the liquids are being transferred. The changes of temperature of the container also need to be taken into consideration for a more accurate result.

Phases (states) of matter and latent heat

DEFINITIONS AND MICROSCOPIC VIEW

When a substance changes phase, the temperature remains constant even though thermal energy is still being transferred.



Cooling curve for molten lead (idealized)

The amount of energy associated with the phase change is called the **latent heat**. The technical term for the change of phase from solid to liquid is **fusion** and the term for the change from liquid to gas is **vaporization**.

The energy given to the molecules does not increase their kinetic energy so it must be increasing their potential energy. Intermolecular bonds are being broken and this takes energy. When the substance freezes bonds are created and this process releases energy.

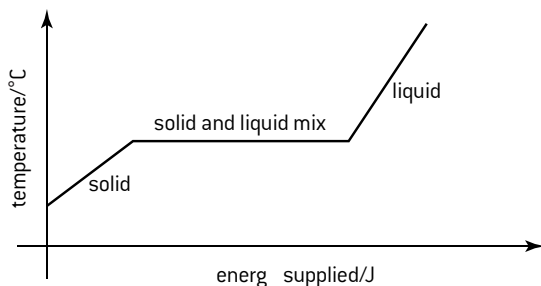
It is a very common mistake to think that the molecules must speed up during a phase change. The molecules in water vapour at 100 °C must be moving with the same average speed as the molecules in liquid water at 100 °C.

The **specific latent heat** of a substance is defined as the amount of energy per unit mass absorbed or released during a change of phase.

In symbols,

$$\text{Specific latent heat } L = \frac{Q}{M} \quad (\text{J kg}^{-1}) \quad Q = ML$$

In the idealized situation of no energy loss, a constant rate of energy transfer into a solid substance would result in a constant rate of increase in temperature until the melting point is reached:



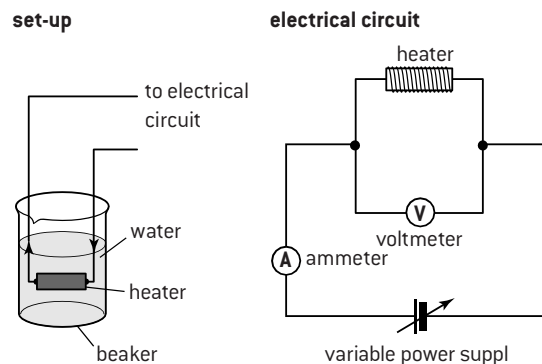
Phase-change graph with temperature vs energy

In the example above, the specific heat capacity of the liquid is less than the specific heat capacity of the solid as the gradient of the line that corresponds to the liquid phase is greater than the gradient of the line that corresponds to the solid phase. A given amount of energy will cause a greater increase in temperature of the liquid when compared with the solid.

METHODS OF MEASURING

The two possible methods of measuring latent heats shown below are very similar in principle to the methods of measuring specific heat capacities (see previous page).

1. A method of measuring the specific latent heat of vaporization of water



The amount of thermal energy provided to water at its boiling point is calculated using electrical energy = $I t V$. The mass vaporized needs to be recorded.

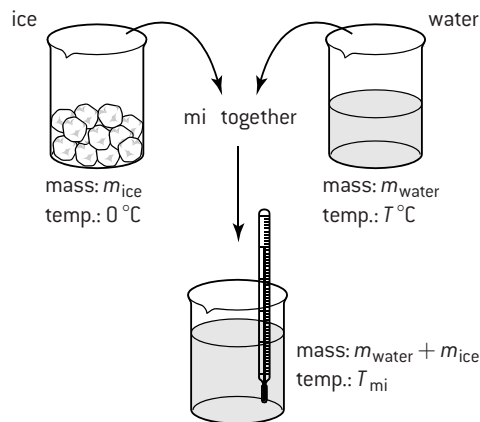
- The specific latent heat $L = \frac{I t V}{(m_1 - m_2)}$.

Sources of experimental error

- Loss of thermal energy from the apparatus.
- Some water vapour will be lost before and after timing.

2. A method of measuring the specific latent heat of fusion of water

Providing we know the specific heat capacity of water, we can calculate the specific latent heat of fusion of water. In the example below, ice (at 0 °C) is added to warm water and the temperature of the resulting mix is measured.



If no energy is lost from the system then,

energy lost by water cooling down = energy gained by ice

$$m_{\text{water}} c_{\text{water}} (T_{\text{water}} - T_{\text{mix}}) = m_{\text{ice}} L_{\text{fusion}} + m_{\text{ice}} c_{\text{water}} T_{\text{mix}}$$

Sources of experimental error

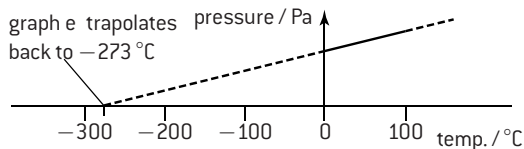
- Loss (or gain) of thermal energy from the apparatus.
- If the ice had not started at exactly zero, then there would be an additional term in the equation in order to account for the energy needed to warm the ice up to 0 °C.
- Water clinging to the ice before the transfer.

The gas laws 1

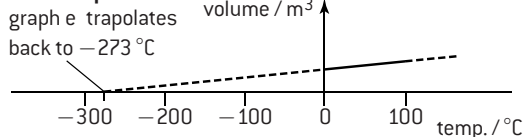
GAS LAWS

For the experimental methods shown below, the graphs below outline what might be observed.

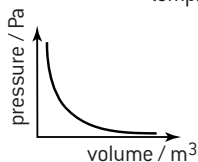
(a) constant volume



(b) constant pressure



(c) constant temperature

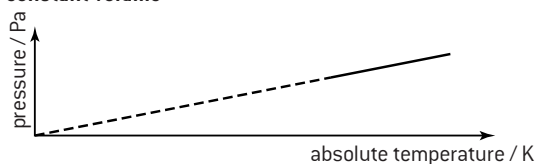


Points to note:

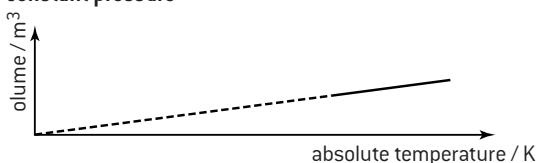
- Although pressure and volume both vary linearly with Celsius temperature, neither pressure nor volume is proportional to Celsius temperature.
- A different sample of gas would produce a different straight-line variation of pressure (or volume) against temperature but both graphs would extrapolate back to the same low temperature, -273°C . This temperature is known as **absolute zero**.
- As pressure increases, the volume decreases. In fact they are inversely proportional.

The trends can be seen more clearly if this information is presented in a slightly different way.

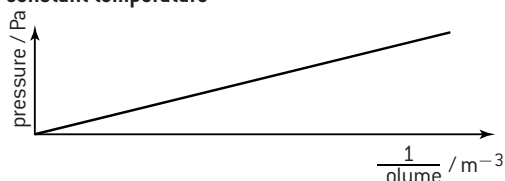
(1) constant volume



(2) constant pressure



(3) constant temperature



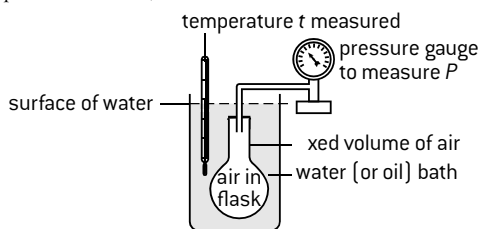
From these graphs for a fixed mass of gas we can say that:

1. At constant V , $p \propto T$ or $\frac{p}{T} = \text{constant}$ (the pressure law)
2. At constant p , $V \propto T$ or $\frac{V}{T} = \text{constant}$ (Charles's law)
3. At constant T , $p \propto \frac{1}{V}$ or $pV = \text{constant}$ (Boyle's law)

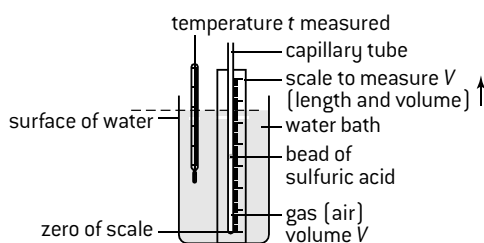
These relationships are known as the **ideal gas laws**. The temperature is always expressed in Kelvin (see page 25). These laws do not always apply to experiments done with real gases. A real gas is said to 'deviate' from ideal behaviour under certain conditions (e.g. high pressure).

EXPERIMENTAL INVESTIGATIONS

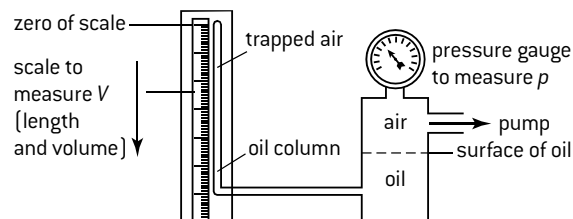
1. Temperature t as the independent variable; P as the dependent variable; V as the control.



- Fixed volume of gas is trapped in the flask. Pressure is measured by a pressure gauge.
 - Temperature of gas altered by temperature of bath – time is needed to ensure bath and gas at same temperature.
2. Temperature t as the independent variable; V as the dependent variable; P as the control.



- Volume of gas is trapped in capillary tube by bead of concentrated sulfuric acid.
 - Concentrated sulfuric acid is used to ensure gas remains dry.
 - Heating gas causes it to expand moving bead.
 - Pressure remains equal to atmospheric.
 - Temperature of gas altered by temperature of bath; time is needed to ensure bath and gas at same temperature.
3. P as the independent variable; V as the dependent variable; t as the control.



- Volume of gas measured against calibrated scale.
- Increase of pressure forces oil column to compress gas.
- Temperature of gas will be altered when volume is changed; time is needed to ensure gas is always at room temperature.

The gas laws 2

EQUATION OF STATE

The three ideal gas laws can be combined together to produce one mathematical relationship.

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

This constant will depend on the mass and type of gas.

If we compare the value of this constant for different masses of different gases, it turns out to depend on the number of molecules that are in the gas – not their type. In this case we use the definition of the mole to state that for n moles of ideal gas

$$\frac{pV}{nT} = \text{a universal constant.}$$

The universal constant is called the **molar gas constant** R .

The SI unit for R is $\text{J mol}^{-1} \text{K}^{-1}$

$$R = 8.314 \text{ J mol}^{-1} \text{K}^{-1}$$

Summary: $\frac{pV}{nT} = R$ Or $pV = nRT$

EXAMPLE

a) What volume will be occupied by 8 g of helium (mass number 4) at room temperature (20°C) and atmospheric pressure ($1.0 \times 10^5 \text{ Pa}$)

$$n = \frac{8}{4} = 2 \text{ moles}$$

$$T = 20 + 273 = 293 \text{ K}$$

$$V = \frac{nRT}{p} = \frac{2 \times 8.314 \times 293}{1.0 \times 10^5} = 0.049 \text{ m}^3$$

b) How many atoms are there in 8 g of helium (mass number 4)?

$$n = \frac{8}{4} = 2 \text{ moles}$$

$$\begin{aligned} \text{number of atoms} &= 2 \times 6.02 \times 10^{23} \\ &= 1.2 \times 10^{24} \end{aligned}$$

DEFINITIONS

The concepts of the **mole**, **molar mass** and the **Avogadro constant** are all introduced so as to be able to relate the mass of a gas (an easily measurable quantity) to the number of molecules that are present in the gas.

Ideal gas

An ideal gas is one that follows the gas laws for all values of p , V and T (see page 29).

Mole

The mole is the basic SI unit of 'amount of substance'. One mole of any substance is equal to the amount of that substance that contains the same number of particles as 0.012 kg of carbon-12 (^{12}C). When writing the unit it is (slightly) shortened to the mol.

Avogadro constant, N_A

This is the number of atoms in 0.012 kg of carbon-12 (^{12}C). It is 6.02×10^{23} .

Molar mass

The mass of one mole of a substance is called the molar mass. A simple rule applies. If an element has a certain mass number, A , then the molar mass will be A grams.

$$n = \frac{N}{N_A}$$

$$\text{number of moles} = \frac{\text{number of atoms}}{\text{Avogadro constant}}$$

IDEAL GASES AND REAL GASES

An ideal gas is one that follows the gas laws for all values of p , V and T and thus ideal gases cannot be liquefied. The microscopic description of an ideal gas is given on page 31. Real gases, however, can approximate to ideal behaviour providing that the intermolecular forces are small enough to be ignored. For this to apply, the pressure/density of the gas must be low and the temperature must be moderate.

LINK BETWEEN THE MACROSCOPIC AND MICROSCOPIC

The equation of state for an ideal gas, $pV = nRT$, links the three macroscopic properties of a gas (p , V and T). Kinetic theory (page 26) describes a gas as being composed of molecules in random motion and for this theory to be valid, each of these macroscopic properties must be linked to the microscopic behaviour of molecules.

A detailed analysis of how a large number of randomly moving molecules interact beautifully predicts another formula that allows the links between the macroscopic and the microscopic to be identified. The derivation of the formula only uses Newton's laws and a handful of assumptions. These assumptions describe from the microscopic perspective what we mean by an ideal gas.

The detail of this derivation is not required by the IB syllabus but the assumptions and the approach are outlined on the following page. The result of this derivation is that the pressure and volume of the idealized gas are related to just two quantities:

$$pV = \frac{2}{3} N \bar{E}_k$$

- The number of molecules present, N
- The average random kinetic energy per molecule, \bar{E}_k .

Equating the right-hand side of this formula with the right-hand side of the macroscopic equation of state for an ideal gas shows that:

$$nRT = \frac{2}{3} N \bar{E}_k$$

But $n = \frac{N}{N_A}$, so

$$\frac{N}{N_A} RT = \frac{2}{3} N \bar{E}_k$$

$$\bar{E}_k = \frac{3}{2} \frac{R}{N_A} T$$

R (the molar gas constant) and N_A (Avogadro constant) are fixed numbers so this equation shows that the absolute temperature is proportional to the average KE per molecule

$$T \propto \bar{E}_k$$

The ratio $\frac{R}{N_A}$ is called the Boltzmann's constant k_B . $k_B = \frac{R}{N_A}$

$$\bar{E}_k = \frac{3}{2} k_B T = \frac{3}{2} \frac{R}{N_A} T$$

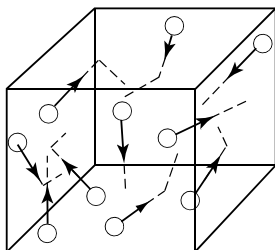
Molecular model of an ideal gas

KINETIC MODEL OF AN IDEAL GAS

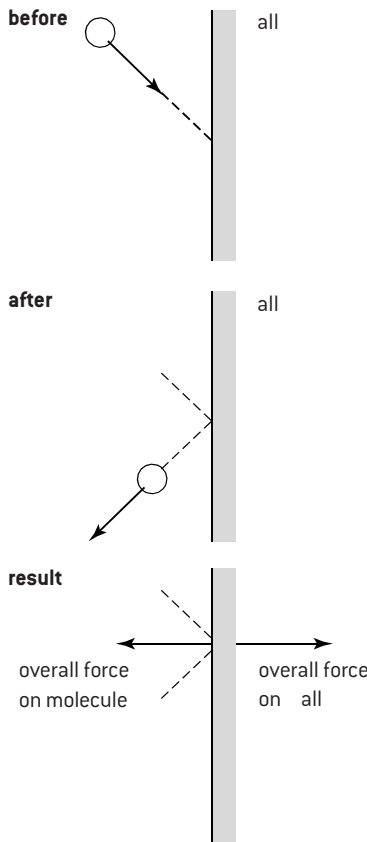
Assumptions:

- Newton's laws apply to molecular behaviour
- there are no intermolecular forces except during a collision
- the molecules are treated as points
- the molecules are in random motion
- the collisions between the molecules are elastic (no energy is lost)
- there is no time spent in these collisions.

The pressure of a gas is explained as follows:



The pressure of a gas is a result of collisions between the molecules and the walls of the container.



A single molecule hitting the walls of the container.

- When a molecule bounces off the walls of a container its momentum changes (due to the change in direction – momentum is a vector).
- There must have been a force on the molecule from the wall (Newton II).
- There must have been an equal and opposite force on the wall from the molecule (Newton III).
- Each time there is a collision between a molecule and the wall, a force is exerted on the wall.
- The average of all the microscopic forces on the wall over a period of time means that there is effectively a constant force on the wall from the gas.
- This force per unit area of the wall is what we call pressure.

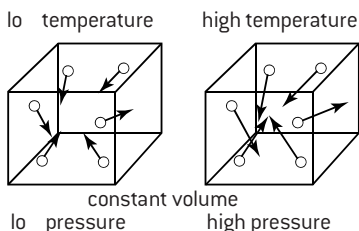
$$P = \frac{F}{A}$$

Since the temperature of a gas is a measure of the average kinetic energy of the molecules, as we lower the temperature of a gas the molecules will move slower. At absolute zero, we imagine the molecules to have zero kinetic energy. We cannot go any lower because we cannot reduce their kinetic energy any further!

PRESSURE LAW

Macroscopically, at a constant volume the pressure of a gas is proportional to its temperature in kelvin (see page 29). Microscopically this can be analysed as follows

- If the temperature of a gas goes up, the molecules have more average kinetic energy – they are moving faster on average.
- Fast moving molecules will have a greater change of momentum when they hit the walls of the container.
- Thus the microscopic force from each molecule will be greater.
- The molecules are moving faster so they hit the walls more often.
- For both these reasons, the total force on the wall goes up.
- Thus the pressure goes up.

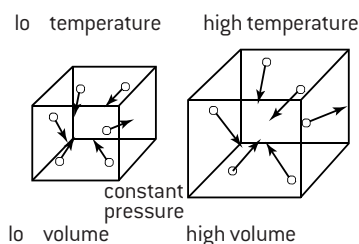


Microscopic justification of the pressure law

CHARLES'S LAW

Macroscopically, at a constant pressure, the volume of a gas is proportional to its temperature in kelvin (see page 29). Microscopically this can be analysed as follows

- A higher temperature means faster moving molecules (see left).
- Faster moving molecules hit the walls with a greater microscopic force (see left).
- If the volume of the gas increases, then the rate at which these collisions take place on a unit area of the wall must go down.
- The average force on a unit area of the wall can thus be the same.
- Thus the pressure remains the same.

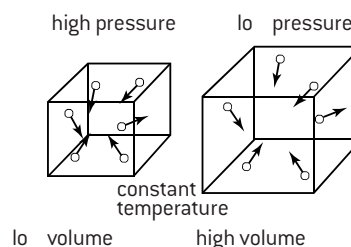


Microscopic justification of Charles's law

BOYLE'S LAW

Macroscopically, at a constant temperature, the pressure of a gas is inversely proportional to its volume (see page 29). Microscopically this can be seen to be correct.

- The constant temperature of gas means that the molecules have a constant average speed.
- The microscopic force that each molecule exerts on the wall will remain constant.
- Increasing the volume of the container decreases the rate with which the molecules hit the wall – average total force decreases.
- If the average total force decreases the pressure decreases.

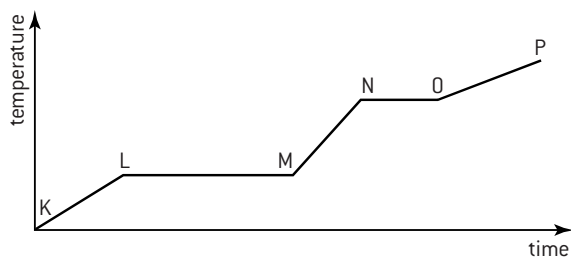


Microscopic justification of Boyle's law

IB Quest ons – thermal phys cs

The following information relates to questions 1 and 2 below.

A substance is heated at a constant rate of energy transfer. A graph of its temperature against time is shown below.



- Which regions of the graph correspond to the substance existing in a mixture of two phases?
 - KL, MN and OP
 - LM and NO
 - All regions
 - No regions
- In which region of the graph is the specific heat capacity of the substance greatest?
 - KL
 - LM
 - MN
 - OP
- When the volume of a gas is isothermally compressed to a smaller volume, the pressure exerted by the gas on the container walls increases. The best microscopic explanation for this pressure increase is that at the smaller volume
 - the individual gas molecules are compressed
 - the gas molecules repel each other more strongly
 - the average velocity of gas molecules hitting the wall is greater
 - the frequency of collisions with gas molecules with the walls is greater
- A lead bullet is fired into an iron plate, where it deforms and stops. As a result, the temperature of the lead increases by an amount ΔT . For an identical bullet hitting the plate with twice the speed, what is the best estimate of the temperature increase?
 - ΔT
 - $2 \Delta T$
 - $2 \Delta T$
 - $4 \Delta T$
- In winter, in some countries, the water in a swimming pool needs to be heated.

- Estimate the cost of heating the water in a typical swimming pool from 5°C to a suitable temperature for swimming. You may choose to consider any reasonable size of pool.

Clearly show any estimated values. The following information will be useful:

Specific heat capacity of water	$4186 \text{ J kg}^{-1} \text{ K}^{-1}$
Density of water	1000 kg m^{-3}
Cost per kW h of electrical energy	$\$0.10$

- Estimated values [4]
- Calculations [7]

- An electrical heater for swimming pools has the following information written on its side:

50 Hz 2.3 kW

- Estimate how many days it would take this heater to heat the water in the swimming pool. [4]
 - Suggest two reasons why this can only be an approximation. [2]
- A cylinder fitted with a piston contains 0.23 mol of helium gas.



The following data are available for the helium with the piston in the position shown.

Volume	$= 5.2 \times 10^{-3} \text{ m}^3$
Pressure	$= 1.0 \times 10^5 \text{ Pa}$
Temperature	$= 290 \text{ K}$

- Use the data to calculate a value for the universal gas constant. (2)
 - State the assumption made in the calculation in (a)(i). (1)
- This question is about determining the specific latent heat of fusion of ice.

A student determines the specific latent heat of fusion of ice at home. She takes some ice from the freezer, measures its mass and mixes it with a known mass of water in an insulating jug. She stirs until all the ice has melted and measures the final temperature of the mixture. She also measured the temperature in the freezer and the initial temperature of the water.

She records her measurements as follows:

Mass of ice used	m_i	0.12 kg
Initial temperature of ice	T_i	-12°C
Initial mass of water	m_w	0.40 kg
Initial temperature of water	T_w	22°C
Final temperature of mixture	T_f	15°C

The specific heat capacities of water and ice are

$$c_w = 4.2 \text{ kJ kg}^{-1} \text{ }^\circ\text{C}^{-1} \text{ and } c_i = 2.1 \text{ kJ kg}^{-1} \text{ }^\circ\text{C}^{-1}$$

- Set up the appropriate equation, representing energy transfers during the process of coming to thermal equilibrium, that will enable her to solve for the specific latent heat L_f of ice. Insert values into the equation from the data above, **but do not solve the equation.** [5]
- Explain the physical meaning of each *energy transfer term* in your equation (but not each symbol). [4]
- State an assumption you have made about the experiment, in setting up your equation in (a). [1]
- Why should she take the temperature of the mixture *immediately* after all the ice has melted? [1]
- Explain from the microscopic point of view, in terms of molecular behaviour, why the temperature of the ice does not increase while it is melting. [4]