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6.5 Ideal Gases

PHYSICS

AQA A Level Revision Notes

A Level Physics AQA

6.5 Ideal Gases

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6.5.1 The Kelvin Scale & Absolute Zero

The Kelvin Scale & Absolute Zero

- On the thermodynamic (Kelvin) temperature scale, absolute zero is defined as:

The lowest temperature possible. Equal to 0 K or –273.15 °C

- It is not possible to have a temperature lower than 0 K
  - This means a temperature in Kelvin will **never** be a negative value
- Absolute zero is defineds as:

The temperature at which the molecules in a substance have zero kinetic energy

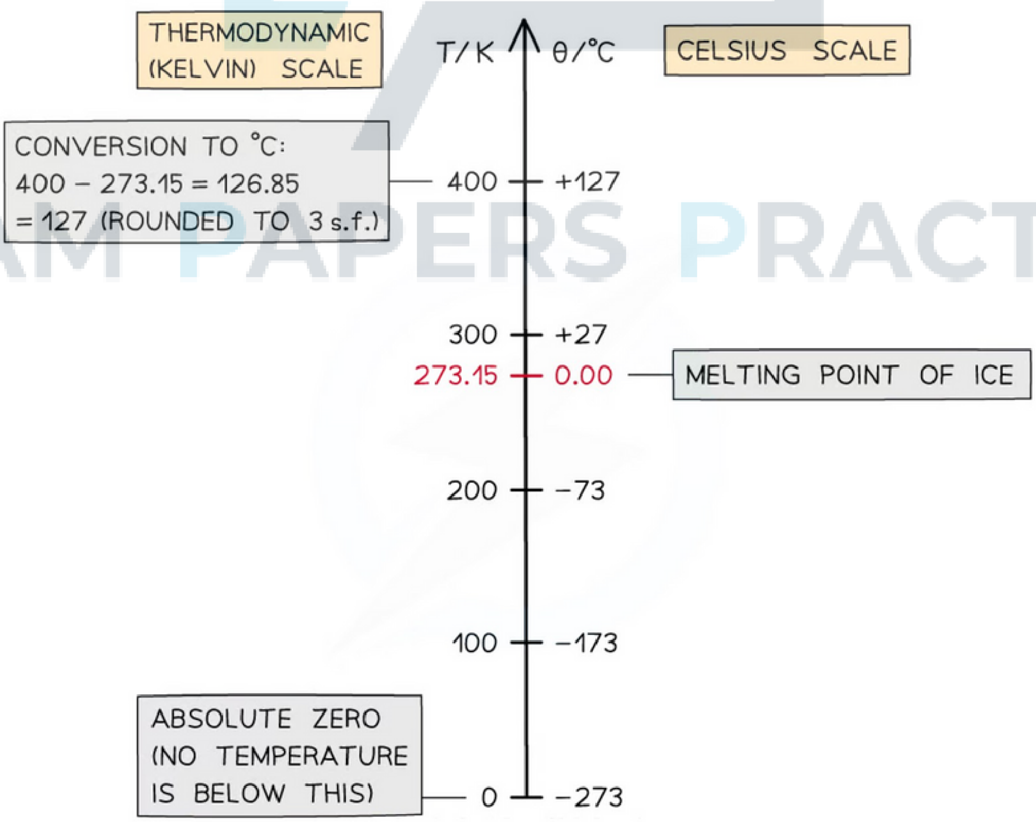
- This means for a system at 0 K, it is not possible to remove any more energy from it
- Even in space, the temperature is roughly 2.7 K, just above absolute zero

Using the Kelvin Scale

- To convert between temperatures θ in the Celsius scale, and T in the Kelvin scale, use the following conversion:

$$\theta / ^\circ C = T / K - 273.15$$

$$T / K = \theta / ^\circ C + 273.15$$



Conversion chart relating the temperature on the Kelvin and Celsius scales

- The divisions on both scales are equal. This means:

A change in a temperature of 1 K is equal to a change in temperature of 1 °C

- This is why when using the specific heat capacity equation

$$Q = mc\Delta\theta$$

- $\Delta\theta$ does not require the temperature to be in either unit
  - This is because the difference in temperature between two values whether in Kelvin or Celsius will be exactly the same

Worked Example

In many ideal gas problems, room temperature is considered to be 300 K. What is this temperature in Celsius?

Step 1: Kelvin to Celsius equation

$$\theta / ^{\circ}C = T / K - 273.15$$

Step 2: Substitute in value of 300 K

$$300\ K - 273.15 = 26.85\ ^{\circ}C$$

Exam Tip

If you forget in the exam whether it's +273.15 or −273.15, just remember that 0 °C = 273.15 K. This way, when you know that you need to +273.15 to a temperature in degrees to get a temperature in Kelvin. For example: $0\ ^{\circ}C + 273.15 = 273.15\ K$.

6.5.2 Gas Laws

Ideal Gas Laws

- The ideal gas laws are the experimental relationships between pressure (*P*), volume (*V*) and temperature (*T*) of an ideal gas
- The mass and the number of molecules of the gas is assumed to be constant for all of these

Boyle's Law

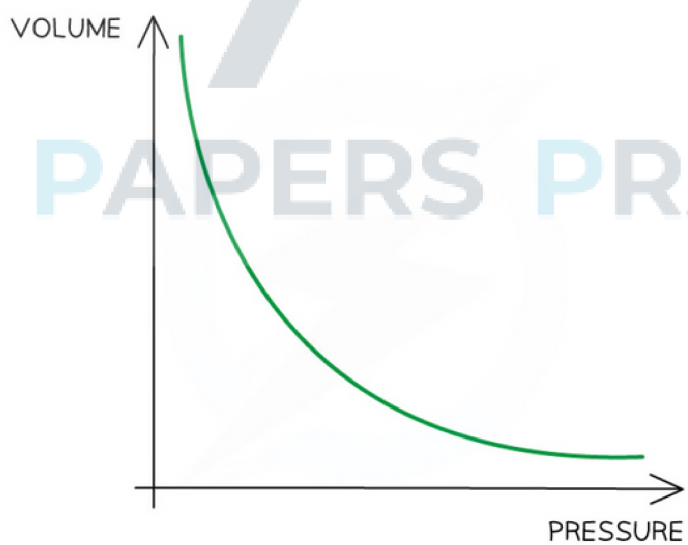
- If the temperature *T* of an ideal gas is constant, then **Boyle's Law** is given by:

$$P \propto \frac{1}{V}$$

- This means the pressure is **inversely proportional** to the volume of a gas
- The relationship between the pressure and volume for a fixed mass of gas at constant temperature can also be written as:

$$P_1V_1 = P_2V_2$$

- Where:
  - P_1 = initial pressure (Pa)
  - P_2 = final pressure (Pa)
  - V_1 = initial volume (m^3)
  - V_2 = final volume (m^3)



Boyle's Law graph representing pressure inversely proportional to volume

- If the temperature increases, the graph is further from the origin and vice versa

Charles's Law

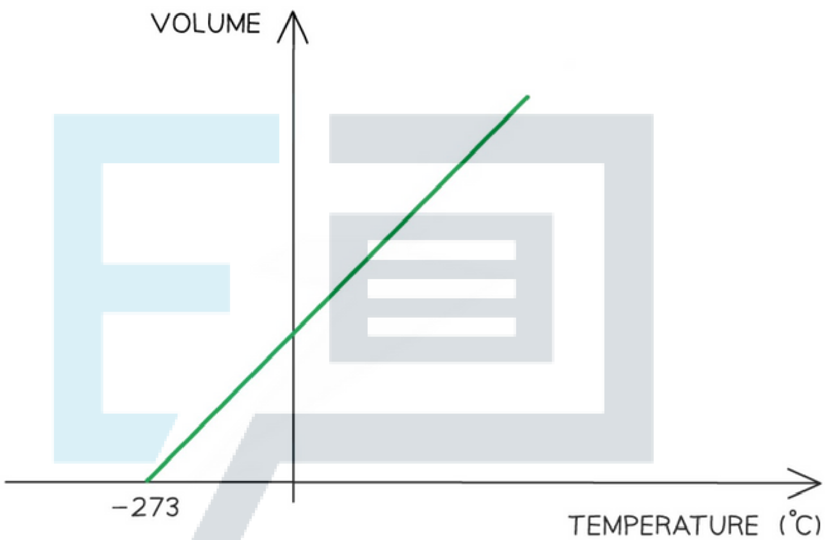
- If the pressure *P* of an ideal gas is constant, then **Charles's law** is given by:

$$V \propto T$$

- This means the volume is **proportional** to the temperature of a gas
- The relationship between the volume and thermodynamic temperature for a fixed mass of gas at constant pressure can also be written as:

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

- Where:
  - V_1 = initial volume (m^3)
  - V_2 = final volume (m^3)
  - T_1 = initial temperature (K)
  - T_2 = final temperature (K)



Charles's Law graph representing temperature (in °C) directly proportional to the volume

- The Charles's Law graph for temperature in **kelvin** against volume is identical except that is a straight line through the **origin**

Pressure Law

- If the volume *V* of an ideal gas is constant, the **Pressure law** is given by:

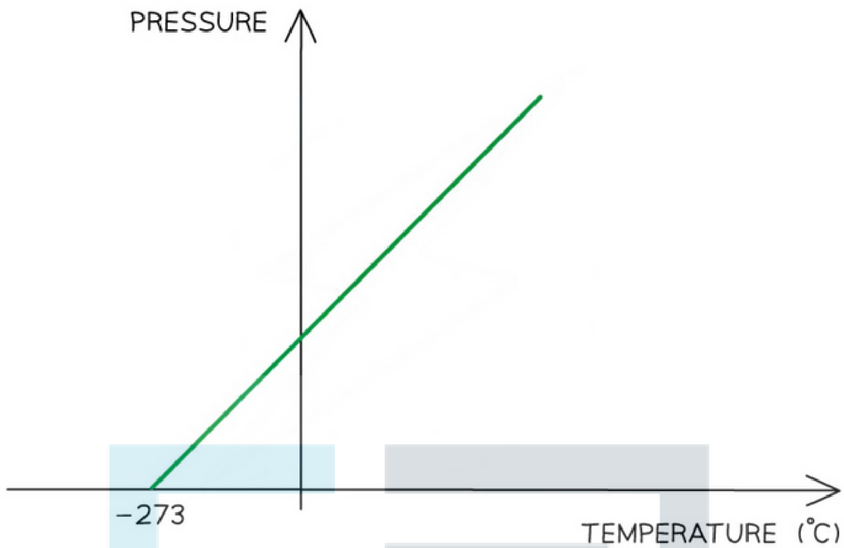
$$P \propto T$$

- This means the pressure is **proportional** to the temperature
- The relationship between the pressure and thermodynamic temperature for a fixed mass of gas at constant volume can also be written as:

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

- Where:
  - P_1 = initial pressure (Pa)
  - P_2 = final pressure (Pa)

- T_1 = initial temperature (K)
- T_2 = final temperature (K)



Pressure Law graph representing temperature (in °C) directly proportional to the volume

Worked Example

The pressure inside a bicycle tyre is 5.10×10^5 Pa when the temperature is 279 K. After the bicycle has been ridden, the temperature of the air in the tyre is 299 K. Calculate the new pressure in the tyre, assuming the volume is unchanged.

Step 1: Choose which ideal gas law to use

Since the volume is constant, the pressure law must be used

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

Step 2: Write down the known quantities

$P_1 = 5.10 \times 10^5$ Pa

$P_2 = ?$

$T_1 = 279$ K

$T_2 = 299$ K

Step 3: Substitute values into pressure law equation

$$P_2 = \frac{P_1 T_2}{T_1} = \frac{(5.10 \times 10^5) \times 299}{279} = 5.47 \times 10^5 \text{ Pa}$$

Exam Tip

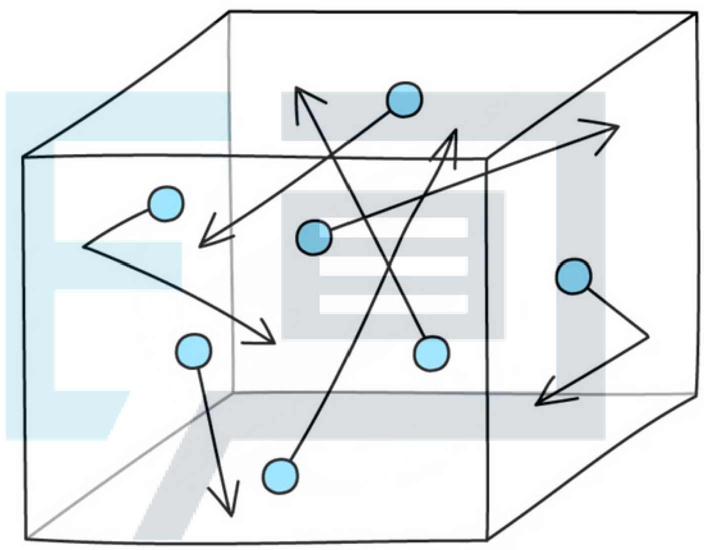
Remember when using any ideal gas law, including the ideal gas equation, the temperature *T* must always be in **kelvin** (K)

Relationships Between Pressure, Volume & Temperature

- An **ideal gas** is one that obeys the relation:

$$pV \propto T$$

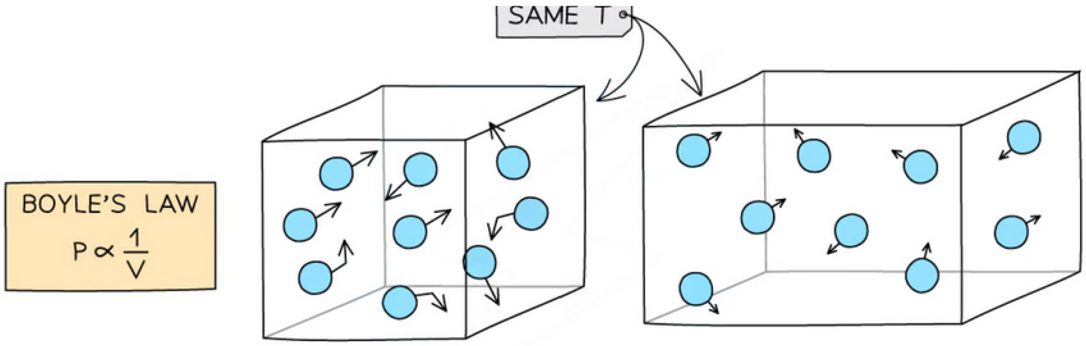
- Where:
  - p = pressure of the gas (Pa)
  - V = volume of the gas (m^3)
  - T = thermodynamic temperature (K)
- The molecules in a gas move around randomly at high speeds, colliding with surfaces and exerting pressure upon them



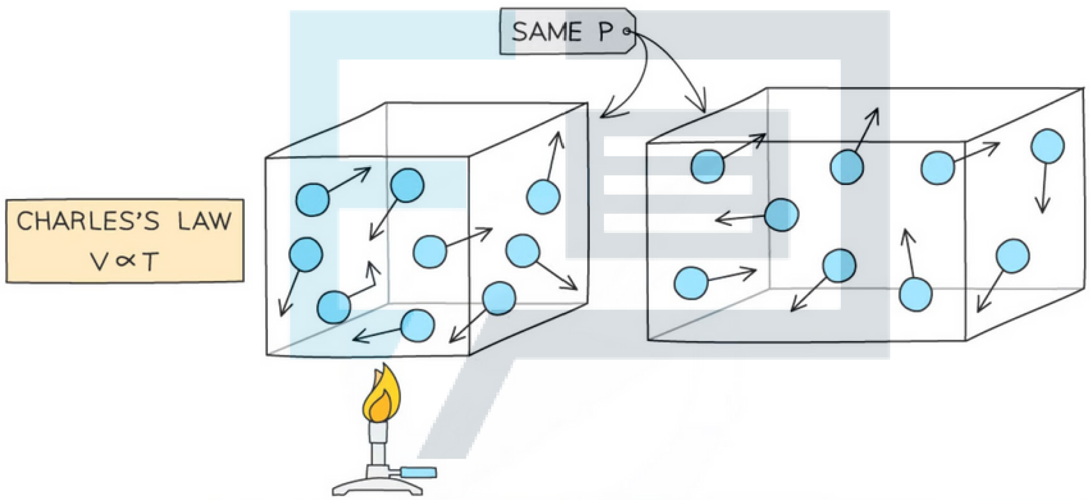
Gas molecules move about randomly at high speeds

- Imagine molecules of gas free to move around in a box
- The temperature of a gas is related to the average speed of the molecules:
  - The hotter the gas, the faster the molecules move
  - Hence the molecules collide with the surface of the walls more frequently
- Since force is the rate of change of momentum:
  - Each collision applies a **force** across the surface area of the walls
  - The faster the molecules hit the walls, the greater the force on them
- Since pressure is the **force per unit area**
  - **Higher temperature leads to higher pressure**
- If the volume V of the box decreases, and the temperature T stays constant:
  - There will be a smaller surface area of the walls and hence more collisions
  - This also creates more pressure
- Since this equates to a greater force per unit area, **pressure** in an ideal gas is therefore defined by:

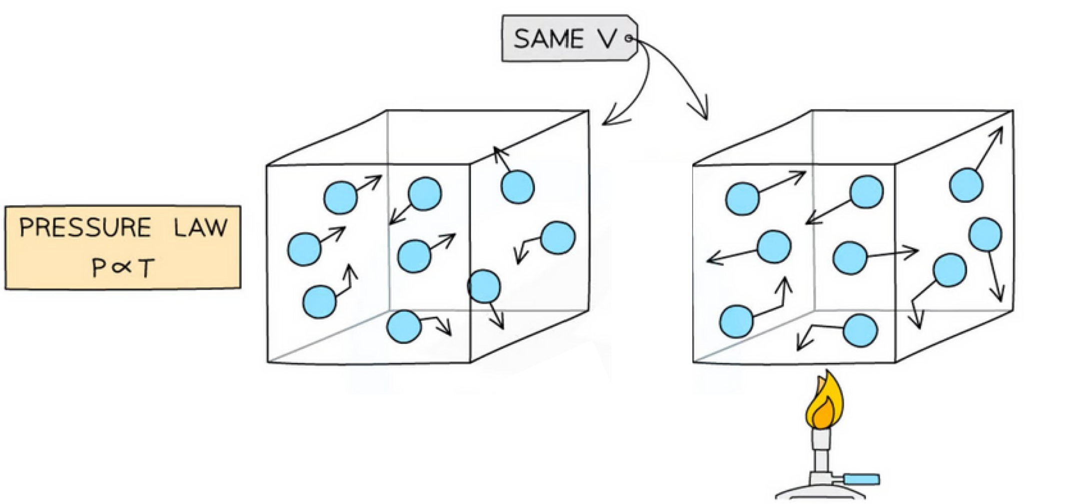
The frequency of collisions of the gas molecules per unit area of a container



IF THE VOLUME OF A GAS IS INCREASED, THE PARTICLES WILL BE FURTHER APART AND WILL COLLIDE LESS WITH EACH OTHER AND THE CONTAINER, DECREASING ITS PRESSURE



IF THE TEMPERATURE OF A GAS IS INCREASED, THE PARTICLES GAIN KINETIC ENERGY AND MOVE FASTER. TO KEEP THE SAME PRESSURE, THEY MUST MOVE FURTHER APART WHICH INCREASES THE VOLUME OF THE GAS



IF THE TEMPERATURE OF A GAS IS INCREASED, THE PARTICLES GAIN KINETIC ENERGY AND MOVE FASTER. THEREFORE THEY WILL COLLIDE MORE WITH EACH OTHER AND THE CONTAINER INCREASING ITS PRESSURE

Molecular model of the three ideal gas laws

Worked Example

An ideal gas is in a container of volume 4.5×10^{-3} m^3. The gas is at a temperature of 30 °C and a pressure of 6.2×10^5 Pa. Calculate the pressure of the ideal gas in the same container when it is heated to 40 °C.

Step 1: Ideal gas relation between pressure, volume and temperature

$$pV \propto T$$

Step 2: Write the equation in full

$$pV = kT$$

- Where k = the constant of proportionality

Step 3: Rearrange for the constant of proportionality

$$k = \frac{pV}{T}$$

Step 4: Convert temperature T into Kelvin

$$\theta \text{ °C} + 273.15 = T \text{ K}$$

$$30 \text{ °C} + 273.15 = 303.15 \text{ K}$$

Step 5: Substitute in known value into constant of proportionality equation

$$k = \frac{6.2 \times 10^5 \times 4.5 \times 10^{-3}}{303.15} = 9.203...$$

Step 6: Rearrange ideal gas relation equation for pressure

$$p = \frac{kT}{V}$$

Step 7: Substitute in new values

$$k = 9.203...$$

$$V \text{ stays the same} = 4.5 \times 10^{-3} \text{ m}^3$$

$$T = 40 \text{ °C} = 40 + 273.15 = 313.15 \text{ K}$$

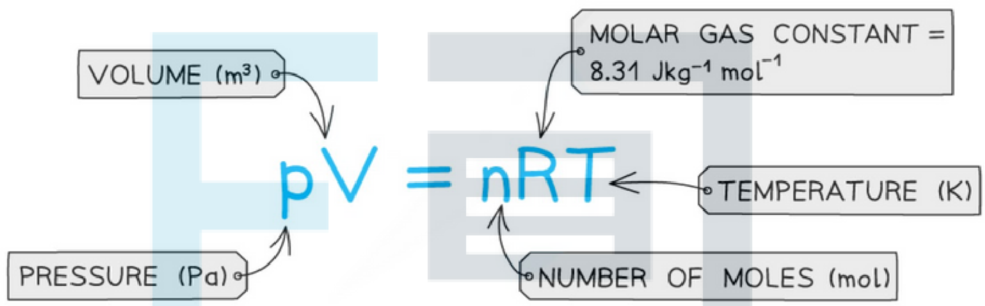
$$p = \frac{(9.203...) \times 313.15}{4.5 \times 10^{-3}} = 640.45... \times 10^3 \text{ Pa} = \mathbf{640\ kPa}$$

Exam Tip

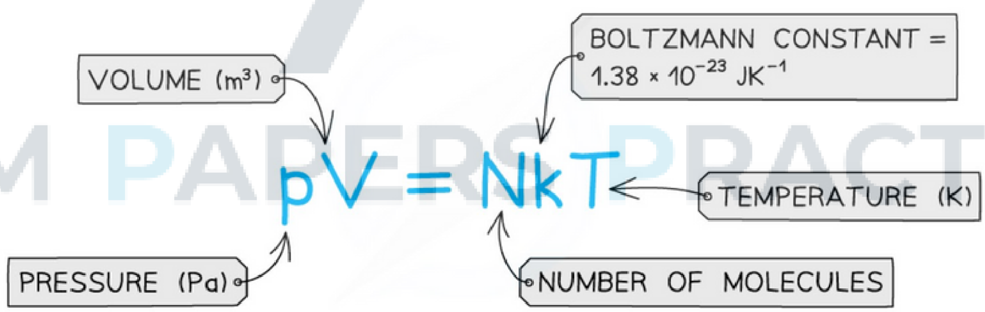
Don't round too early in your working out! In the worked example, the unrounded value of *k* is represented by "..." to show its full value is to be carried over to the next step of the calculation. On your calculator, this can be done by using the "ans" button instead of typing in the whole number.

Ideal Gas Equation

- An ideal gas is a specific type of gas which:
  - Has molecules with negligible volume
  - Collisions which are elastic
  - Cannot be liquified
  - Has no interactions between the molecules (except during collisions)
  - Obeys the (ideal) gas laws (Boyles law, Charles' law and Pressure law)
- All of these can occur at any temperature or pressure
- The equation of state for an ideal gas (or the ideal gas equation) can be expressed as:



- The ideal gas equation can also be written in the form:



- An ideal gas is therefore defined as:

A gas which obeys the equation of state pV = nRT at all pressures, volumes and temperatures

? Worked Example

A storage cylinder of an ideal gas has a volume of 8.3×10^3 cm^3. The gas is at a temperature of 15 °C and a pressure of 4.5×10^7 Pa. Calculate the amount of gas in the cylinder, in moles.

Step 1: Write down the ideal gas equation

Since the number of moles (n) is required, use the equation:

$$pV = nRT$$

Step 2: Rearrange for the number of moles n

$$n = \frac{pV}{RT}$$

Step 3: Substitute in values

$$V = 8.3 \times 10^3\ cm^3 = 8.3 \times 10^3 \times 10^{-6} = 8.3 \times 10^{-3}\ m^3$$

$$T = 15\ °C + 273.15 = 288.15\ K$$

$$n = \frac{4.5 \times 10^7 \times 8.3 \times 10^{-3}}{8.31 \times 288.15} = 155.98 = \mathbf{160\ mol}\ (2\ s.f.)$$

Exam Tip

Don't worry about remembering the values of R and k, they will both be given in the equation sheet in your exam.

6.5.4 Work Done by a Gas

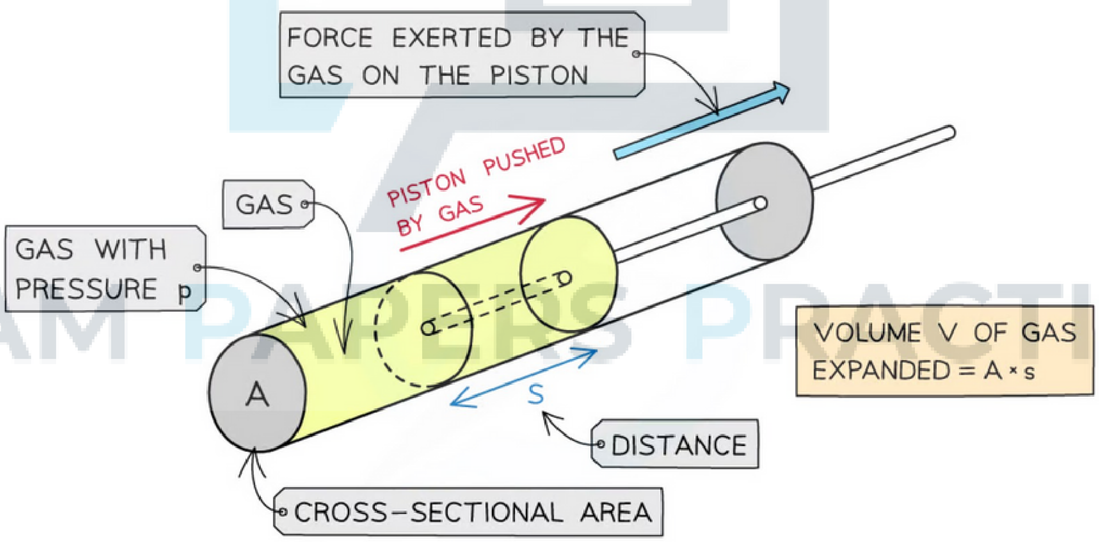
Work Done by a Gas

- When a gas expands, it does work on its surroundings by exerting pressure on the walls of the container it's in
- This is important, for example, in a steam engine where expanding steam pushes a piston to turn the engine
- The work done when a volume of gas changes at constant pressure is defined as:

$$W = p\Delta V$$

- Where:
  - W = work done (J)
  - p = external pressure (Pa)
  - V = volume of gas (m^3)

- For a gas inside a cylinder enclosed by a moveable piston, the force exerted by the gas pushes the piston outwards
- Therefore, the gas **does work on the piston**



The gas expansion pushes the piston a distance s

Derivation

- The volume of gas is at constant pressure. This means the force *F* exerted by the gas on the piston is equal to :

$$F = p \times A$$

- Where:
  - *p* = pressure of the gas (Pa)
  - *A* = cross-sectional area of the cylinder (m^2)

- The definition of work done is:

$$W = F \times s$$

- Where:
  - F = force (N)
  - s = displacement in the direction of force (m)
- The displacement of the gas d multiplied by the cross-sectional area A is the increase in volume ΔV of the gas:

$$W = p \times A \times s$$

- This gives the equation for the work done when the volume of a gas changes at constant pressure:

$$W = p\Delta V$$

- Where:
  - ΔV = increase in the volume of the gas in the piston when expanding (m^3)

- This is assuming that the surrounding pressure p does not change as the gas expands
- This will be true if the gas is expanding against the pressure of the atmosphere, which changes very slowly
- When the gas **expands** (V increases), work is done **by** the gas
- When the gas is **compressed** (V decreases), work is done **on** the gas

Worked Example

When a balloon is inflated, its rubber walls push against the air around it. Calculate the work done when the balloon is blown up from 0.015 m^3 to 0.030 m^3. Atmospheric pressure = 1.0×10^5 Pa.

Step 1: Write down the equation for the work done by a gas

$$W = p\Delta V$$

Step 2: Substitute in values

$$\Delta V = \text{final volume} - \text{initial volume} = 0.030 - 0.015 = 0.015 \text{ m}^3$$

$$W = (1.0 \times 10^5) \times 0.015 = 1500 \text{ J}$$

Exam Tip

The pressure p in the work done by a gas equation is not the pressure of the gas but the pressure of the surroundings. This is because when a gas expands, it does work **on** the surroundings.

Avogadro, Molar Gas & Boltzmann Constant

Avogadro's Constant

- The atomic mass unit (u) is approximately the mass of a proton or neutron = 1.66×10^{-27} kg
- This means that an atom or molecule has a mass approximately equal to the number of protons and neutrons it contains
- A carbon-12 atom has a mass of:

$$\mathbf{12\,u = 12 \times 1.66 \times 10^{-27} = 1.99 \times 10^{-26}\,kg}$$

- The exact number for a mole is defined as the number of molecules in exactly 12 g of carbon:

$$1 \text{ mole} = \frac{0.012}{1.99 \times 10^{-26}} = 6.02 \times 10^{23} \text{ molecules}$$

- Avogadro's constant (N_A) is defined as:

The number of atoms of carbon-12 in 12 g of carbon-12; equal to 6.02×10^{23} mol^{-1}

- For example, 1 mole of sodium (Na) contains 6.02×10^{23} **atoms** of sodium
- The number of atoms can be determined if the number of moles is known by multiplying by N_A, for example:

2.0 mol of nitrogen contains: $2.0 \times N_A = 2.0 \times 6.02 \times 10^{23} = 1.20 \times 10^{24}$ atoms

Moles and Atomic Mass

- One mole of any element is equal to the **relative atomic mass** of that element in grams
  - For example, helium has an atomic mass of 4, meaning 1 mole of helium has a mass of 4 g

- If the substance is a compound, add up the relative atomic masses, for example, water (H_2O) is made up of
  - 2 hydrogen atoms (each with an atomic mass of 1) and 1 oxygen atom (atomic mass of 16)
  - So, 1 mole of water would have a mass of (2 × 1) + 16 = 18 g

Molar Mass

- The molar mass of a substance is the mass, in grams, in one mole
  - Its unit is **g mol^{-1}**

- The number of moles from this can be calculated using the equation:

$$\text{Number of moles} = \frac{\text{mass (g)}}{\text{molar mass (g mol}^{-1})}$$

Boltzmann & The Molar Gas Constant

- The Boltzmann constant k is used in the ideal gas equation and is defined by the equation:

$$k = \frac{R}{N_A}$$

- Where:
  - R = molar gas constant
  - N_A = Avogadro's constant
- Boltzmann's constant, therefore, has a value of:

$$k = \frac{8.31}{6.02 \times 10^{23}} = 1.38 \times 10^{-23} \text{ J K}^{-1}$$

- The Boltzmann constant relates the properties of microscopic particles (e.g. kinetic energy of gas molecules) to their macroscopic properties (e.g. temperature)
  - This is why the units are **J K^{-1}**
- Its value is very small because the increase in kinetic energy of a molecule is very small for every incremental increase in temperature

Worked Example

How many molecules are there in 6 g of magnesium-24?

Step 1: Calculate the mass of 1 mole of magnesium

One mole of any element is equal to the relative atomic mass of that element in grams

1 mole = 24 g of magnesium

Step 2: Calculate the amount of moles in 6 g

$$\frac{6}{24} = 0.25 \text{ moles}$$

Step 3: Convert the moles to number of molecules

1 mole = 6.02×10^{23} molecules

0.25 moles = $0.25 \times 6.02 \times 10^{23} = 1.51 \times 10^{23}$ molecules

Exam Tip

If you want to find out more about the mole, check out the AQA A Level Chemistry revision notes!

Required Practical: Investigating Gas Laws

Investigating Boyle's Law

- The overall aim of this experiment is to investigate the effect of Boyle's Law
  - This is the effect of pressure on volume at a constant temperature
- This is just one example of how this required practical might be tackled

Variables

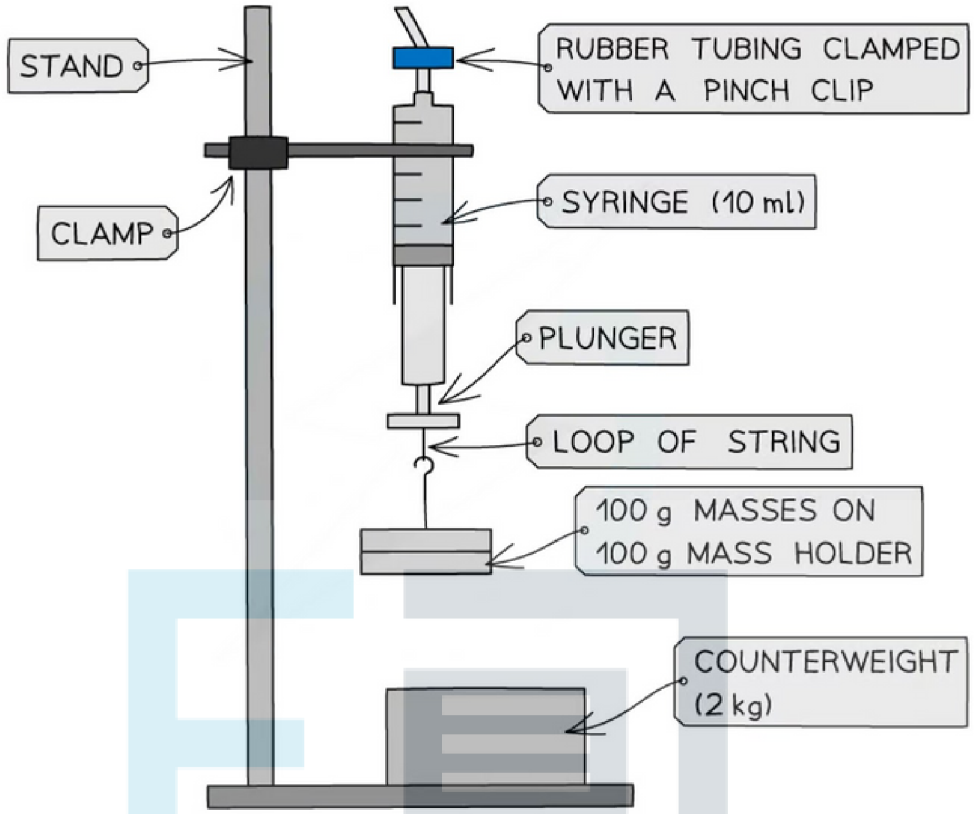
- Independent variable = Mass, *m* (kg)
- Dependent variable = Volume, *V* (m^3)
- Control variables:
  - Temperature
  - Cross-sectional area of the syringe

Equipment List

Apparatus	Purpose
Clamp Stand	To hold the equipment
Syringe (10 ml)	To measure the volume of the air inside the tube
Rubber Tubing	To stop air getting into the syringe
Pinch Clip	To hold the rubber tubing in place
String	To hang the masses
100g masses with 100g holder	Used to change the volume of the air
Counterweight (or G clamp)	To keep the stand and clamp secure and rigid on the flat surface
Vernier Caliper	To calculate the diameter of the syringe

- Resolution of measuring equipment:
  - Pressure gauge = 0.02×10^5 Pa
  - Volume = 0.2 cm^3
  - Vernier Caliper = 0.02 mm

Method



Apparatus setup for Boyle's Law

1. With the plunger removed from the syringe, measure the inside diameter, *d* of the syringe using a vernier calliper. Remember to take at least 3 repeat readings and find an average
2. The plunger should be replaced and the rubber tubing should be fit over the nozzle and clamped with a pinch clip as close to the nozzle as possible
3. Set up the apparatus as shown in the diagram and make sure the temperature of the room will remain constant throughout
4. Push the syringe upwards until it reads the lowest volume of air visible. Record this volume
5. Add the 100 g mass holder with a 100 g mass on it to the loop of string at the bottom of the plunger. Wait a few seconds to ensure the temperature is kept constant since work is done against the plunger when the volume increases
6. Record the value of the new volume from the syringe scale
7. Repeat the experiment by adding two 100 g masses at a time up to 8–10 readings. This is so a significant change in volume can be seen each time
8. Record the mass and volume

- An example table of results might look like this:

READ FROM THE SYRINGE SCALE (→ VOLUME / cm³)

NO MASS HOLDER (→ 0.0)

100g MASS HOLDER AND A 100g MASS (→ 0.2)

MASS / kg	VOLUME / cm^3
0.0	
0.2	
0.4	
0.6	
0.8	
1.0	
1.2	

DIAMETER OF THE SYRINGE	DIAMETER 1/mm	DIAMETER 2/mm	DIAMETER 3/mm	AVERAGE DIAMETER/mm

Analysing the Results

- Boyle's Law can be represented by the equation:

$$pV = \text{constant}$$

- This means the pressure must be calculated from the experiment
- The exerted pressure of the masses is calculated by:

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

- Where:
  - F = weight of the masses, mg (N)
  - A = cross-sectional area of the syringe (m^2)
- The cross-sectional area is found from the equation for the area of a circle:

$$A = \frac{\pi d^2}{4}$$

- To calculate the pressure of the gas:

Pressure of the gas = Atmospheric pressure – Exerted pressure from the masses

- Where:
  - Atmospheric pressure = 101 kPa

- The table of results may need to be modified to fit these extra calculations. Here is an example of how this might look:

$F = mg$ → FORCE / N

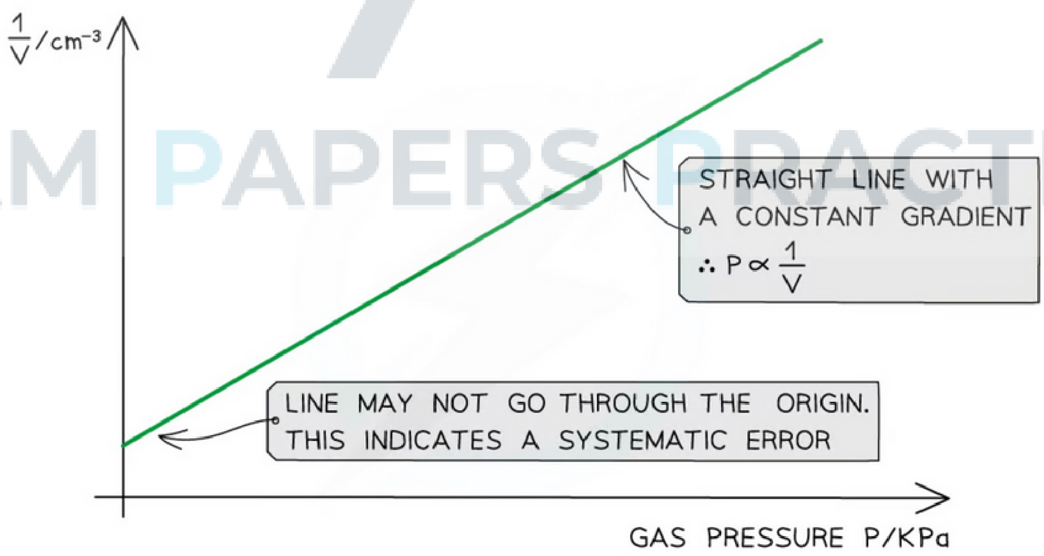
$P = \frac{F}{A}$ → EXERTED PRESSURE / Pa

ATMOSPHERIC PRESSURE − EXERTED PRESSURE → GAS PRESSURE P/KPa

MASS / kg	FORCE / N	EXERTED PRESSURE / Pa	GAS PRESSURE P/KPa	VOLUME V/cm^3	$\frac{1}{V}$/cm^{-3}
0.0					
0.2					
0.4					
0.6					
0.8					
1.0					
1.2					

- Once these values are calculated:

1. Plot a graph of *p* against 1 / *V* and draw a line of best fit
2. If this plot is a straight line graph, this means that the pressure is proportional to the inverse of the volume, hence confirming Boyle's Law (pV = constant)



Evaluating the Experiment

Systematic Errors:

- There may be friction in the syringe which causes a systematic error
  - Use a syringe that has very little friction or lubricated it, so the only force is from the weights pulling the syringe downwards

Random Errors:

- The reading of the volume should be taken a few seconds after the mass has been added to the holder
  - Otherwise, a reading will be taken when the temperature is not constant
- This experiment is prone to many random errors with the equipment and surrounding temperature
  - Make sure to take repeat readings to decrease the effect of these

Safety Considerations

- A counterweight or G-clamp must be used to avoid the stand toppling over and causing injury, especially if the surface is not completely flat

Investigating Charles's Law

- The overall aim of this experiment is to investigate the effects of Charles's law, which is the effect of volume on temperature at constant pressure
- This is just one example of how this required practical might be tackled

Variables

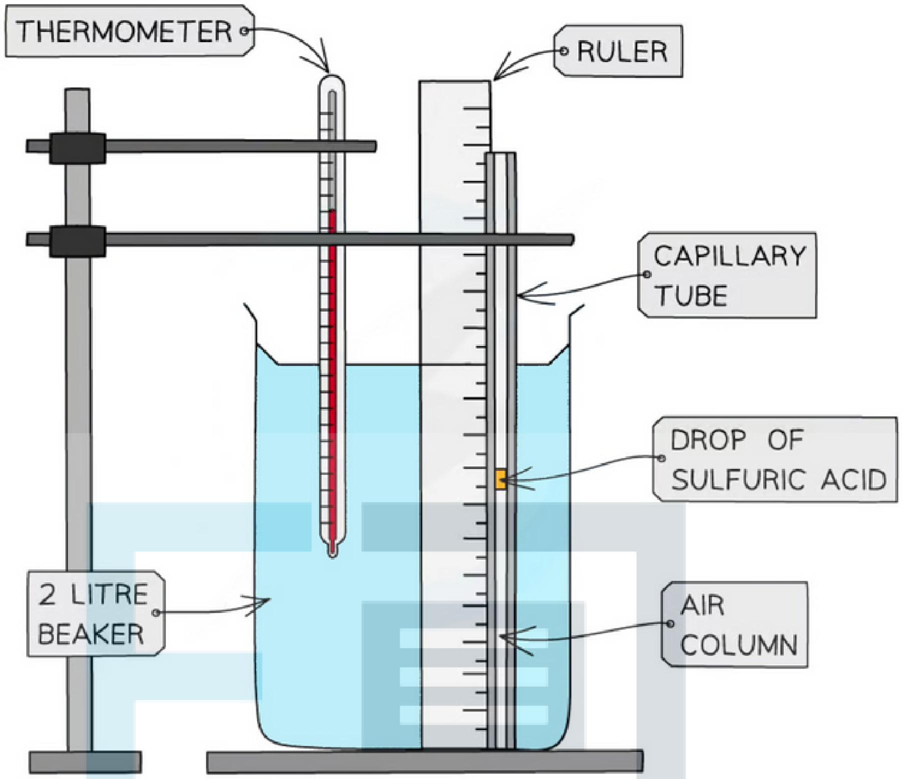
- Independent variable = Temperature, *T* (°C)
- Dependent variable = Height of the gas, *h* (cm)
- Control variables:
  - Pressure

Equipment List

Apparatus	Purpose
Clamp Stand	To hold the equipment
Capillary Tube (atleast 25 cm long)	To hold the gas (e.g. dry air)
Concentrated Sulfuric Acid	To trap the gas in the capillary tube
2 Litre Beaker	To hold the capillary tube, thermometer and the water
30 cm Ruler	To measure the height of the capillary tube
Thermometer	To measure the temperature of the water and the gas
Kettle	To boil water

- Resolution of measuring equipment:
  - 30 cm ruler = 1 mm
  - 2 litre beaker = 50 ml

Method



Apparatus setup for Charles's Law

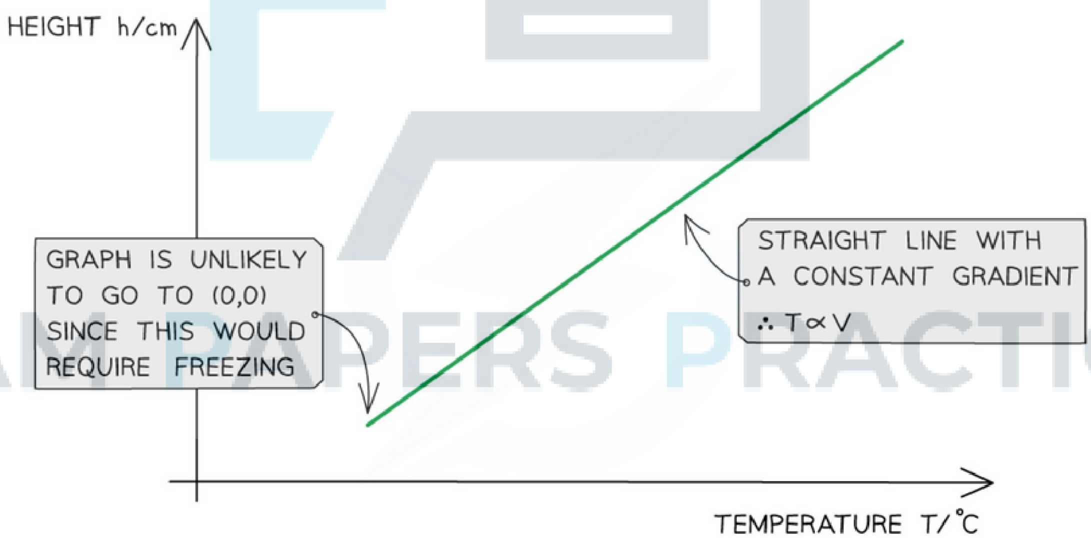
1. The capillary tube should have one open end at the top and a closed end at the bottom. This is to keep the pressure at atmospheric pressure, and constant. Assume the temperature of the water is the same as the temperature of the gas
2. Set up the apparatus as shown in the diagram. Make sure the drop of sulfuric acid is halfway up the tube
3. Boil some water in a kettle and pour it into the beaker for the full 2 litres. Make sure the water covers all the gas, and stir well
4. When the temperature goes down to 95 °C, read the height of the gas from the ruler Make sure this value is read from eye level on the ruler
5. Record the height of the gas as the temperature decreases in increments of 5 °C. Make sure you have at least 8 readings or down to room temperature

- An example table of results might look like:

TEMPERATURE T/°C	HEIGHT h/cm
95	
90	
85	
80	
75	
70	
65	
60	

Analysing the Results

1. Plot a graph of the height of the gas in cm and the temperature in °C
2. Draw a line of best fit
3. Calculate the gradient



4. If this is a straight-line graph, then this means the temperature is proportional to the height. Since the height is proportional to the volume ($V = \pi r^2 h$) then this means Charles's law is confirmed, and the temperature is proportional to the volume too

- To find a value of absolute zero T_0, the equation of the graph can be written as

$$\mathbf{h = mT + c}$$

- Comparing this to the equation of a straight line: y = mx + c
  - y = h
  - x = T
  - m = gradient
  - c = y-intercept

1. Plot a graph of the height of the gas in cm and the temperature in °C
2. Draw a line of best fit
3. Calculate the gradient

- Since c is a constant:

$$h_0 - mT_0 = h_1 - mT_1$$

- At absolute zero, $h_0 = 0$

$$-mT_0 = h_1 - mT_1$$

$$T_0 = \frac{h_1 - mT_1}{-m} = T_1 - \frac{h_1}{m}$$

- Picking any co-ordinate of h and t from the line of best fit, and substituting into the equation will give a value of absolute zero
- Check this value is close to the accepted value of –273°C

Evaluating the Experiment

Systematic Errors:

- Make sure the capillary tube is close to the ruler and properly aligned to get an accurate value of the height of the gas
  - Otherwise, the reading taken will be slightly out each time

Random Errors:

- Although this is a slower process, the experiment can be repeated by measuring the height as the gas cools instead
- There can be parallax error when taking the temperature and height readings by reading them at eye level
- Stir the water well so it is the same temperature throughout the beaker, and so the gas is the same temperature as well

Safety Considerations

- When using boiling water, make sure not to spill it onto your skin or any electrical equipment
- Make sure the bench is protected with a heat-proof mat so the boiling water does not damage the surface

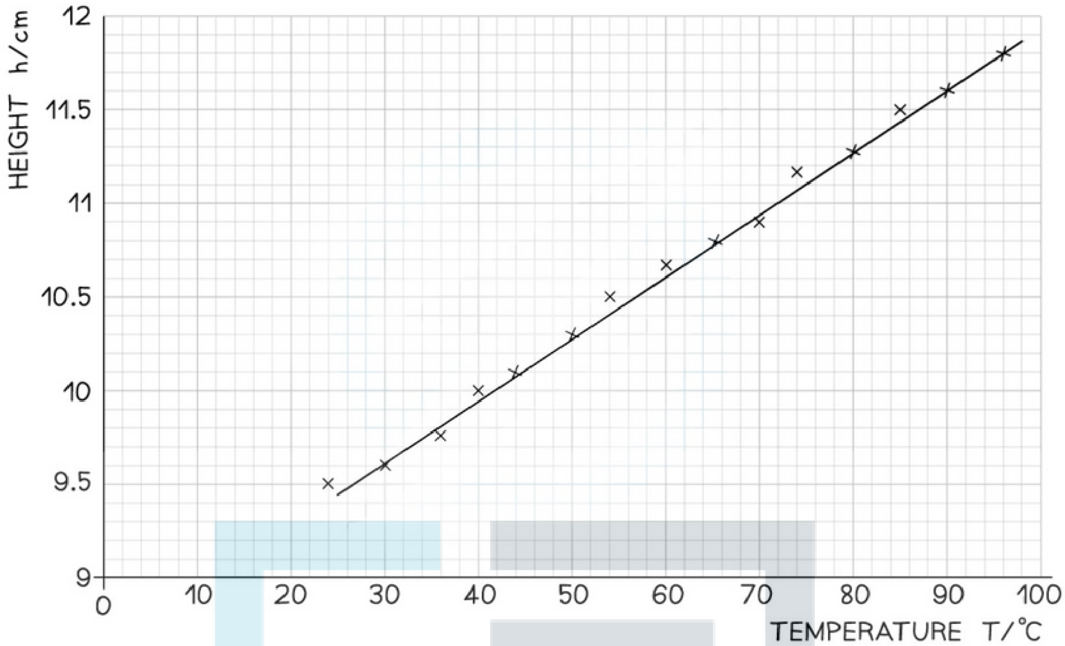
Worked Example

A student investigates the relationship between the temperature and volume of a column of air. They obtain the following results:

Temperature/°C	Height / cm
95	11.8
90	11.6
85	11.5
80	11.3
75	11.2
70	10.9
65	10.8
60	10.7
55	10.5
50	10.3
45	10.1
40	10.0
35	9.8
30	9.6
25	9.5

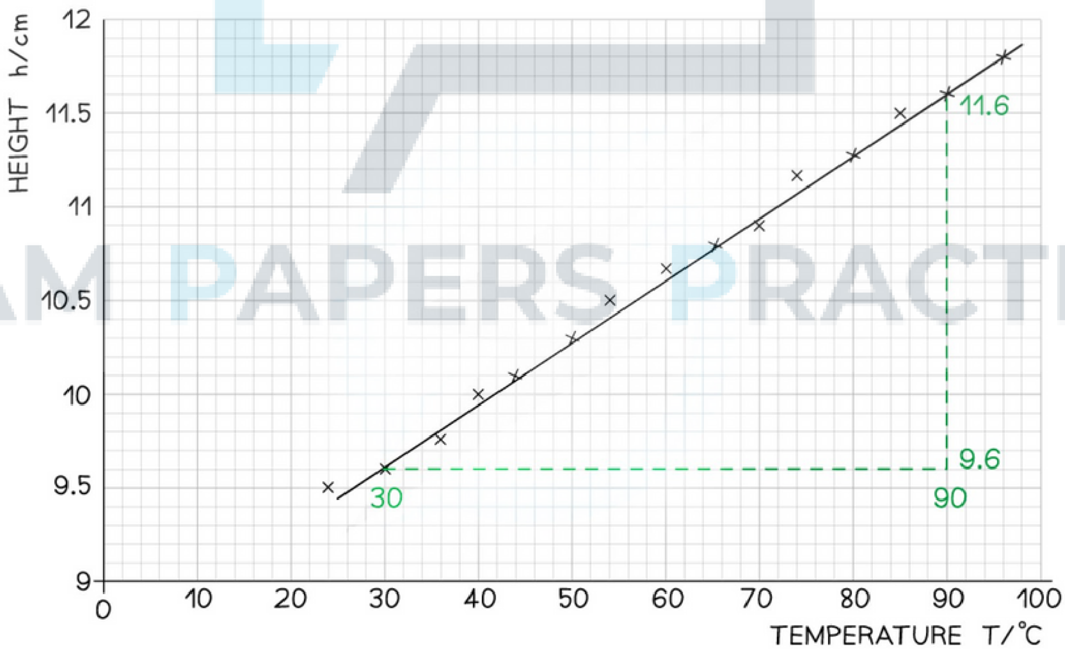
Calculate the value of absolute zero from these results and its relative percentage error with the accepted value of −273.15 °C

Step 1: Plot a graph of temperature *T* against volume *V*



- Make sure the axes are properly labelled and the line of best fit is drawn with a ruler

Step 2: Calculate the gradient of the graph



- The gradient is calculated by:

$$\text{gradient} = \frac{11.6 - 9.6}{90 - 30} = 0.033 \text{ cm / °C}$$

Step 3: Calculate the value of absolute zero

$$T_0 = \frac{h_1 - mT_1}{-m} = T_1 - \frac{h_1}{m}$$

- Where T_0 is absolute zero and (T_1, h_1) is any co-ordinate on the line of best fit
- Using the coordinates (60, 10.6):

$$T_0 = 60 - \frac{10.6}{0.033} = -261.2121 = -261 \text{ °C}$$

Step 4: Calculate its relative percentage error with the accepted value of –273.15 °C

$$\text{Relative percentage error} = \frac{\text{Measured value} - \text{Accepted Value}}{\text{Accepted Value}} \times 100\ \%$$

$$\text{Relative percentage error} = \frac{-261.2121 - (-273.15)}{-273.15} \times 100\ \% = -4.37\ \%$$