CIE International A Level Physics

Topic 14, 15 & 16

Thermal Physics

Revision Booklet

This booklet covers:

- Temperature and Thermal Equilibrium
- Temperature Scales
- Specific Heat Capacity and Latent Heat
- The Mole and Ideal Gases
- Kinetic Theory of Gases
- Internal Energy
- The First Law of Thermodynamics

Thermal Equilibrium

Thermal Energy Transfer

Thermal (heat) energy is transferred from a region of **higher temperature** to a region of **lower temperature**. Transfer continues until the two regions reach the same temperature.

Thermal Equilibrium

Two objects are in **thermal equilibrium** when there is no net transfer of thermal energy between them. This occurs when they are at the **same temperature**.

Zeroth Law of Thermodynamics: If object A is in thermal equilibrium with object C, and object B is also in thermal equilibrium with object C, then A and B are in thermal equilibrium with each other.

Temperature Scales

Thermometric Property

A **thermometric property** is a physical property that varies continuously and measurably with temperature. Examples include:

- Resistance of a metal wire (e.g. platinum resistance thermometer)
- e.m.f. of a thermocouple
- Volume of a gas at constant pressure
- Density of a liquid

Thermodynamic Temperature Scale

The **thermodynamic (Kelvin) temperature scale** does not depend on the property of any particular substance. It is an absolute scale with:

- **Absolute zero** (0 K) — the lowest possible temperature
- **Triple point of water** (273.16 K) — fixed reference point

Temperature Conversion

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

T = thermodynamic temperature (K)
θ = Celsius temperature (°C)

Note: a temperature *difference* of 1 K equals a difference of 1 °C.

Absolute Zero

At absolute zero (0 K = −273.15 °C):

- Molecules have minimum possible kinetic energy
- It is impossible to reach in practice, only to approach asymptotically
- All ideal gases would have zero volume (zero pressure at constant volume)

Specific Heat Capacity and Latent Heat

Specific Heat Capacity

The **specific heat capacity** c of a substance is the energy required to raise the temperature of **unit mass** by **one kelvin** (or one degree Celsius), without change of state.

$$c = \frac{Q}{m\,\Delta T} \qquad \text{units: J kg}^{-1}\text{K}^{-1}$$

Specific Heat Capacity

$$Q = mc\Delta T$$

Q = thermal energy transferred (J)
m = mass (kg)
c = specific heat capacity (J kg^{-1} K^{-1})
ΔT = temperature change (K or °C)

Specific Latent Heat

The **specific latent heat** L of a substance is the energy required to change the state of **unit mass** at constant temperature.

$$L = \frac{Q}{m} \qquad \text{units: J kg}^{-1}$$

- **Specific latent heat of fusion** L_f: solid → liquid (melting)
- **Specific latent heat of vaporisation** L_v: liquid → gas (boiling)

Note: $L_v > L_f$ for any given substance, since greater work is done separating molecules completely during vaporisation.

Specific Latent Heat

$$Q = mL$$

Q = thermal energy transferred (J)
m = mass (kg)
L = specific latent heat (J kg^{-1})

Common Mistake

During a change of state, temperature remains **constant** even though energy is being supplied. The energy goes into breaking intermolecular bonds, not increasing kinetic energy.

The Mole and Ideal Gases

Amount of Substance

The **mole** (mol) is the SI base unit for amount of substance. One mole of any substance contains exactly $N_A = 6.02 \times 10^{23}$ particles (the Avogadro constant).

Equation of State for an Ideal Gas

$$pV = nRT \qquad \text{or} \qquad pV = NkT$$

p = pressure (Pa)
V = volume (m^3)
n = number of moles (mol)
R = molar gas constant = 8.31 $\text{J mol}^{-1}\text{K}^{-1}$
T = thermodynamic temperature (K)
N = number of molecules
k = Boltzmann constant = 1.38×10^{-23} J K^{-1}

The Boltzmann constant: $k = R/N_A$

Ideal Gas

An **ideal gas** is one that obeys $pV \propto T$ (where T is thermodynamic temperature) under all conditions. Real gases approximate ideal behaviour at low pressures and high temperatures.

Gas Laws — special cases of $pV = nRT$

- **Boyle's Law** (constant T): pV = constant, so $p_1V_1 = p_2V_2$
- **Charles' Law** (constant p): V/T = constant, so $V_1/T_1 = V_2/T_2$
- **Pressure Law** (constant V): p/T = constant, so $p_1/T_1 = p_2/T_2$

Common Mistake

Always use **thermodynamic temperature in kelvin** in gas law calculations — never degrees Celsius.

Kinetic Theory of Gases

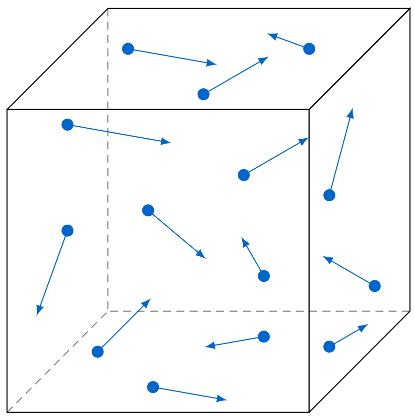
Assumptions of the Kinetic Model

Basic Assumptions of Kinetic Theory

1. The gas contains a **large number** of molecules moving in **random directions** with a range of speeds.
2. The molecules occupy **negligible volume** compared to the volume of the gas.
3. All collisions are **perfectly elastic** (no kinetic energy lost).
4. **Intermolecular forces are negligible** except during collisions.
5. The duration of collisions is **negligible** compared to the time between collisions.
6. Molecules obey **Newton's laws of motion**.

Pressure from Kinetic Theory

The pressure exerted by a gas arises from molecules colliding with the container walls. Each collision exerts a force; the average of many such collisions gives a steady pressure.



Kinetic Theory Equation

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

p = pressure (Pa)
V = volume (m^3)
N = total number of molecules
m = mass of one molecule (kg)
$\langle c^2\rangle$ = mean square speed ($m^2\,s^{-2}$)

The **root-mean-square speed**: $c_{r.m.s.} = \sqrt{\langle c^2\rangle}$

Note: $c_{r.m.s.}$ is not the same as the mean speed $\bar{c}$. Because we square first then average, faster molecules contribute more. In general $c_{r.m.s.} > \bar{c}$.

Derivation of $pV = \frac{1}{3}Nm\langle c^2\rangle$

Consider one molecule of mass m moving with speed c_x in the x-direction inside a box of side L.

Step 1 — change in momentum at one wall: The molecule hits the right wall and bounces back elastically:

$$\Delta p = mc_x - (-mc_x) = 2mc_x$$

Step 2 — time between collisions with the same wall: The molecule must travel $2L$ before returning:

$$\Delta t = \frac{2L}{c_x}$$

Step 3 — force from one molecule:

$$F = \frac{\Delta p}{\Delta t} = \frac{2mc_x}{2L/c_x} = \frac{mc_x^2}{L}$$

Step 4 — pressure from N molecules: Summing over all N molecules and dividing by area L^2:

$$p = \frac{Nm\langle c_x^2\rangle}{L^3} = \frac{Nm\langle c_x^2\rangle}{V}$$

Step 5 — extend to 3 dimensions: By symmetry of random motion: $\langle c_x^2\rangle = \langle c_y^2\rangle = \langle c_z^2\rangle = \frac{1}{3}\langle c^2\rangle$

$$\boxed{pV = \tfrac{1}{3}Nm\langle c^2\rangle}$$

Molecular Kinetic Energy

Derivation of $\langle E_k\rangle = \frac{3}{2}kT$

Start with the two equations for an ideal gas:

$$pV = \tfrac{1}{3}Nm\langle c^2\rangle \qquad \text{and} \qquad pV = NkT$$

Equating the right-hand sides:

$$\tfrac{1}{3}Nm\langle c^2\rangle = NkT$$

Divide both sides by N and multiply by $\frac{3}{2}$:

$$\tfrac{1}{2}m\langle c^2\rangle = \tfrac{3}{2}kT$$

Since $\langle E_k\rangle = \frac{1}{2}m\langle c^2\rangle$:

$$\boxed{\langle E_k\rangle = \tfrac{3}{2}kT}$$

Average Translational Kinetic Energy

$$\langle E_k\rangle = \frac{1}{2}m\langle c^2\rangle = \frac{3}{2}kT$$

The average kinetic energy of a molecule is **directly proportional to thermodynamic temperature**.

Key Deductions

- At absolute zero, molecules have minimum kinetic energy
- Molecules of different gases at the same temperature have the **same average KE**
- Heavier molecules move more **slowly** on average at the same temperature: $c_{\text{r.m.s.}} = \sqrt{3kT/m}$
- $c_{\text{r.m.s.}} \propto \sqrt{T}$ — doubling thermodynamic temperature increases r.m.s. speed by factor $\sqrt{2}$

Internal Energy

Internal Energy

The **internal energy** of a system is the sum of the **random kinetic energies** and **potential energies** of all the molecules in the system.

$$U = E_{k,\text{total}} + E_{p,\text{total}}$$

- Internal energy is a **function of state**.
- For an **ideal gas**: $E_p = 0$, so $U = E_{k,\text{total}}$ only.
- A rise in temperature ⇒ increase in average KE ⇒ increase in U.
- During a change of state: temperature (and KE) constant; E_p increases as bonds break; U increases.

The First Law of Thermodynamics

Work Done by/on a Gas

When a gas changes volume at **constant pressure**:

$$W = p\Delta V$$

W = work done **by** the gas (J)
p = pressure (Pa)
ΔV = change in volume (m^3)

First Law of Thermodynamics

$$\Delta U = q + W$$

ΔU = increase in internal energy of the system (J)
q = energy supplied **to** the system by heating (J)
W = work done **on** the system (J)

Sign Convention

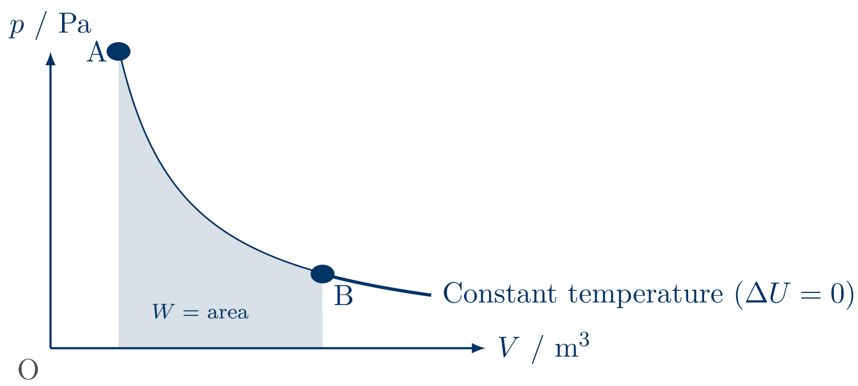
CIE uses $\Delta U = q + W$ where W is work done **on** the system. Some textbooks use $\Delta U = q - W$ where W is work done **by** the system. Always check which convention is being used.

Special Cases of the First Law

Process	Condition	Result
Constant temperature	$\Delta T = 0 \Rightarrow \Delta U = 0$	$q = -W$
No heat transfer	$q = 0$	$\Delta U = W$
Constant volume	$\Delta V = 0 \Rightarrow W = 0$	$\Delta U = q$
Constant pressure	p constant	$\Delta U = q - p\Delta V$

p–V Diagrams

The **area under a p–V curve** equals the work done by the gas during that process.



Formula Summary

Formula	Quantity	Units
$T/\text{K} = \theta/°\text{C} + 273.15$	Temperature conversion	K
$Q = mc\Delta T$	Sensible heat	J
$Q = mL$	Latent heat	J
$pV = nRT$	Ideal gas (moles)	Pa, m^3, K
$pV = NkT$	Ideal gas (molecules)	Pa, m^3, K
$k = R/N_A$	Boltzmann constant	$\text{J}\,\text{K}^{-1}$
$pV = \frac{1}{3}Nm\langle c^2\rangle$	Kinetic theory	Pa
$\langle E_k\rangle = \frac{3}{2}kT$	Mean molecular KE	J
$W = p\Delta V$	Work done by gas	J
$\Delta U = q + W$	First law	J

Constants: $R = 8.31\ \text{J}\,\text{mol}^{-1}\text{K}^{-1}$ $k = 1.38 \times 10^{-23}\ \text{J}\,\text{K}^{-1}$ $N_A = 6.02 \times 10^{23}\ \text{mol}^{-1}$

Worked Examples

Example 1 — Specific Heat Capacity

Question: A 2.0 kg block of aluminium ($c = 900\ \text{J}\,\text{kg}^{-1}\text{K}^{-1}$) is heated from 20 °C to 80 °C. Calculate the energy supplied.

Solution

$$Q = mc\Delta T = 2.0 \times 900 \times (80 - 20) = \mathbf{1.08 \times 10^5}\ \text{J}$$

Example 2 — Ideal Gas Law

Question: A fixed mass of ideal gas has pressure 1.2×10^5 Pa and volume 3.0×10^{-3} m^3 at 27 °C. It is heated at constant pressure until its volume doubles. Find the final temperature.

Solution

$T_1 = 27 + 273.15 = 300.15$ K. At constant pressure, $V \propto T$:

$$T_2 = T_1 \times \frac{V_2}{V_1} = 300.15 \times 2 = \mathbf{600}\ \text{K}\ (\mathbf{327}\ °\text{C})$$

Example 3 — r.m.s. Speed

Question: Calculate the r.m.s. speed of nitrogen molecules ($M = 0.028\ \text{kg mol}^{-1}$) at 300 K.

Solution

Mass of one molecule: $m = 0.028/6.02 \times 10^{23} = 4.65 \times 10^{-26}$ kg

$$c_{\text{r.m.s.}} = \sqrt{\frac{3kT}{m}} = \sqrt{\frac{3 \times 1.38 \times 10^{-23} \times 300}{4.65 \times 10^{-26}}} = \sqrt{2.67 \times 10^{5}} = \mathbf{517}\ \text{m s}^{-1}$$

Example 4 — First Law

Question: A gas is compressed. 650 J of work is done on the gas and 200 J of heat flows out. Find the change in internal energy.

Solution

$W = +650$ J (work done *on* gas), $q = -200$ J (heat leaves)

$$\Delta U = q + W = -200 + 650 = +\mathbf{450}\ \text{J}$$

Internal energy **increases** by 450 J.

Practice Exam Questions

Section A — Short Answer Questions

Q1. Define *specific latent heat of vaporisation* and explain why it is greater than the specific latent heat of fusion for the same substance.

[3 marks]

Q2. State the basic assumptions of the kinetic theory of gases.

[4 marks]

Q3. Explain what is meant by *internal energy* and state how it differs for an ideal gas compared to a real gas.

[3 marks]

Q4. A gas at pressure 2.0×10^5 Pa and temperature 300 K occupies a volume of 0.40 m^3. Calculate the number of molecules present.

[2 marks]

Q5. Show that the mean kinetic energy of a gas molecule is $\frac{3}{2}kT$ by combining the kinetic theory equation with the ideal gas equation.

[3 marks]

Section B — Longer Structured Questions

Q6. A sample of water of mass 0.50 kg is heated from 20 °C to 100 °C and then completely vaporised.
($c_{\text{water}} = 4200\ \text{J kg}^{-1}\text{K}^{-1}$, $L_v = 2.26 \times 10^6\ \text{J kg}^{-1}$)

(a) Calculate the energy needed to heat the water from 20 °C to 100 °C.

[2 marks]

(b) Calculate the energy needed to vaporise the water at 100 °C.

[2 marks]

(c) Explain in terms of molecular behaviour why energy is needed to vaporise water even though the temperature does not change.

[2 marks]

Q7. A fixed mass of ideal gas undergoes the cycle A → B → C → A on a p–V diagram where A → B is constant temperature expansion, B → C is constant volume pressure decrease, and C → A is constant pressure compression.

(a) State the change in internal energy during A → B and justify your answer. *[2 marks]*

(b) Using the first law, find q for A → B given that the work done by the gas is 440 J. *[2 marks]*

(c) Describe the energy changes during the constant volume process B → C. *[2 marks]*

Mark Scheme and Answers

Q1. Specific latent heat of vaporisation is the energy per unit mass required to change a substance from liquid to gas at constant temperature [1]; greater than L_f because molecules must be completely separated, requiring work against intermolecular forces over a much greater distance than in melting [2].

Q2. Large number of molecules moving randomly in all directions with a range of speeds [1]; occupying negligible volume compared to container [1]; collisions perfectly elastic [1]; intermolecular forces negligible except during collisions; collisions of negligible duration; obey Newton's laws [1].

Q3. Internal energy is the sum of random kinetic and potential energies of all molecules [1]; for an ideal gas, $E_p = 0$ so internal energy is kinetic energy only [1]; for a real gas, molecules have potential energy due to intermolecular forces [1].

Q4. $N = pV/kT = (2.0 \times 10^5 \times 0.40)/(1.38 \times 10^{-23} \times 300) = \mathbf{1.93 \times 10^{25}}$ molecules [2].

Q5. $pV = \frac{1}{3}Nm\langle c^2\rangle$ [1]; equate with $pV = NkT$: $\frac{1}{3}m\langle c^2\rangle = kT$ [1]; so $\frac{1}{2}m\langle c^2\rangle = \frac{3}{2}kT$ [1].

Q6(a). $Q = 0.50 \times 4200 \times 80 = \mathbf{1.68 \times 10^5}$ J [2].

Q6(b). $Q = 0.50 \times 2.26 \times 10^6 = \mathbf{1.13 \times 10^6}$ J [2].

Q6(c). KE of molecules unchanged (temperature constant) [1]; energy breaks intermolecular bonds as molecules separate completely, increasing potential energy [1].

Q7(a). $\Delta U = 0$ [1]; temperature is constant and for an ideal gas internal energy depends only on temperature [1].

Q7(b). $\Delta U = 0$, work done on gas $= -440$ J; $q = +440$ J — heat flows into gas [2].

Q7(c). No work done ($\Delta V = 0$) [1]; temperature and KE decrease; heat flows out; internal energy decreases [1].

Revision Checklist

	Learning Objective	Confidence (1–3)
☐	Explain thermal equilibrium and direction of heat flow	
☐	Convert between Celsius and Kelvin	
☐	State examples of thermometric properties	
☐	Define and use specific heat capacity: $Q = mc\Delta T$	
☐	Define and use specific latent heat: $Q = mL$	
☐	State the assumptions of kinetic theory	
☐	Use $pV = nRT$ and $pV = NkT$	
☐	Derive and use $pV = \frac{1}{3}Nm\langle c^2\rangle$	
☐	Show that mean KE of a molecule is $\frac{3}{2}kT$	
☐	Define internal energy and relate to temperature	
☐	Apply the first law $\Delta U = q + W$	
☐	Interpret and use p–V diagrams	

Key: 1 = Need more work   2 = Getting there   3 = Confident

Good luck with your revision!

Remember: understanding *why* formulas work is more powerful than memorising them.
Practice drawing p–V diagrams and deriving key results from first principles.