Cambridge AS & A Level · Physics 9702 · Notes

Thermodynamics A2

Read online at **pastopic.com/cambridge/physics-9702/notes/thermodynamics**

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Read first: **Temperature**, **Ideal gases**

What you need to know

By the end of this topic you should be able to:

1. **Say what internal energy is.** The internal energy of a system is the sum of the kinetic and potential energies of its molecules, which are spread randomly among them. It depends only on the **state** of the system (its pressure, volume and temperature), not on how it got there.
2. **Link temperature and internal energy.** A rise in the temperature of an object means its molecules have more kinetic energy on average, so its internal energy has increased.
3. **Recall and use** $W = p\Delta V$ for the work done when a gas changes volume at constant pressure, and tell apart the work done **by** the gas (when it expands) and the work done **on** the gas (when it is compressed).
4. **Recall and use the first law of thermodynamics**, $\Delta U = q + W$: the increase in internal energy equals the thermal energy supplied to the system by heating plus the work done on the system.

You need the ideas of energy from IGCSE, and from earlier A Level topics: the ideal gas equation and $U = \frac{3}{2}NkT$ for an ideal gas (Ideal gases), and specific heat capacity and specific latent heat (Temperature). Exam questions on this topic often use them.

Understand it

Internal energy

Every object is made of molecules that move about at random and pull or push on each other.

- Because they move, the molecules have **kinetic energy**. At a higher temperature they move faster on average, so the total kinetic energy is larger.
- Because of the forces between them, the molecules have **potential energy**, which depends on how far apart they are. Pulling molecules further apart against the attractive forces between them increases their potential energy (like lifting a mass away from the Earth).

The **internal energy** U is the **sum of the random kinetic and potential energies of all the molecules**. The energies are shared out at random: at any moment some molecules are fast and some are slow. The kinetic energy of the object moving as a whole, or its gravitational potential energy, is **not** part of its internal energy.

U is a property of the state. A sample of gas with a given pressure, volume and temperature always has the same internal energy, whatever happened to it before. So if a gas is taken round a **complete cycle** back to its starting state, its internal energy is back where it started: $\Delta U = 0$ for the cycle.

Temperature and internal energy. The mean kinetic energy of the molecules depends on the temperature. So when the temperature of an object rises, the kinetic energy of its molecules rises, and its internal energy increases.

Which energy changes? In any process, ask two questions: does the temperature change (kinetic energy), and does the separation of the molecules change (potential energy)?

Process	Kinetic energy	Potential energy	Internal energy
Gas heated at constant volume	increases (temperature rises)	no change (same separation)	increases
Water boiling at $100\,^{\circ}$C	no change (temperature constant)	increases (molecules pulled apart)	increases
Ice melting at $0\,^{\circ}$C	no change	increases	increases
Wire stretched at constant temperature	no change	increases (separation increases)	increases
Steam condensing at $100\,^{\circ}$C	no change	decreases	decreases

During a change of state the temperature stays constant even though energy is being supplied: the energy goes into **potential energy** (pulling molecules apart), not kinetic energy.

An ideal gas has no potential energy. In an ideal gas there are no forces between the molecules, so their potential energy is zero. Its internal energy is just the total kinetic energy of its molecules:

$$U = N \times \tfrac{3}{2}kT = \tfrac{3}{2}NkT = \tfrac{3}{2}nRT = \tfrac{3}{2}pV$$

So for an ideal gas $U \propto T$, and $U \propto pV$: on a pressure–volume graph the internal energy is larger wherever pV is larger. A change at **constant temperature** has $\Delta U = 0$.

Two ways to change the internal energy

There are exactly two ways to change the internal energy of a system:

- **Heating (or cooling).** Thermal energy flows into the system from something hotter (or out of it to something colder). Call the energy **supplied to** the system by heating q. If energy leaves by cooling, q is negative.
- **Doing work.** A force moves something. For a gas, this means a change in volume. Call the work done **on** the system W. If the system does work on its surroundings, W is negative.

Heat and work are **transfers** of energy, not energy the system "has": a gas does not contain heat or work, it contains internal energy.

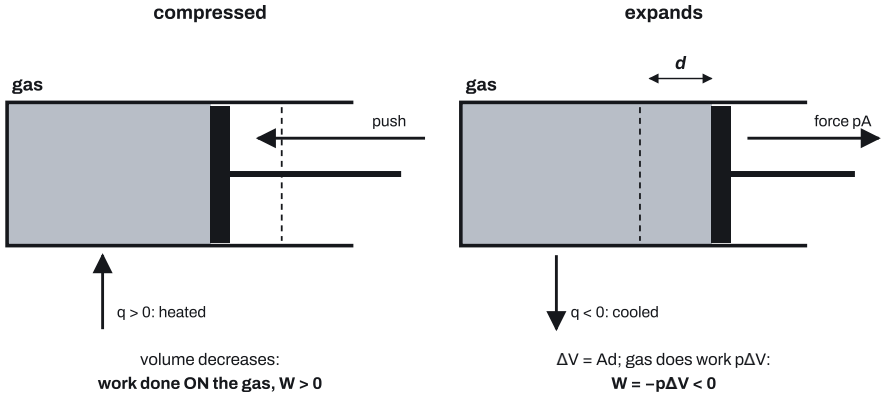
The first law of thermodynamics

Energy is conserved, so the increase in internal energy is the energy that went in by both routes:

$$\Delta U = q + W$$

- ΔU = **increase** in internal energy of the system;
- q = thermal energy **supplied to** the system by heating;
- W = work done **on** the system.

Each quantity is counted as energy going **into** the system. Energy going out gives a negative value. "ΔU" is already a change: say "U increases" or "ΔU is positive", not "ΔU increases".



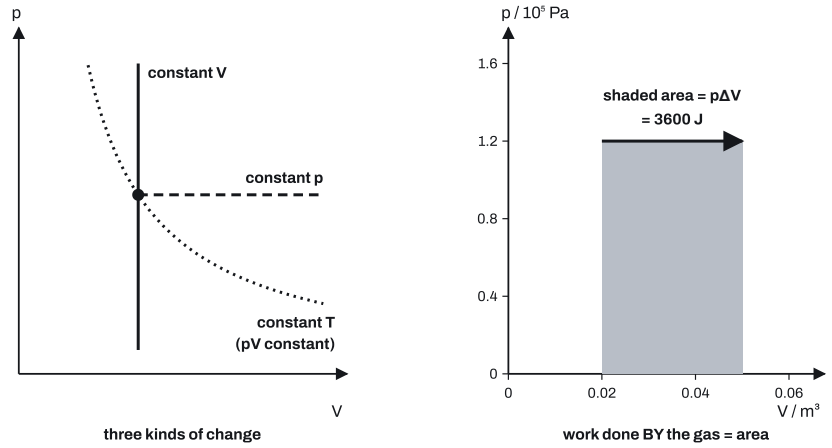
Work done when a gas changes volume

A gas at pressure p pushes on a piston of area A with a force $F = pA$. If the piston moves out a distance d while the pressure stays constant, the gas does work

$$\text{work done by the gas} = Fd = pAd = p\Delta V$$

because Ad is the increase in volume ΔV.

- **Expansion** (ΔV positive): the gas does work on its surroundings, so the work done **on** the gas is **negative**: $W = -p\Delta V$.
- **Compression** (ΔV negative): the surroundings do work on the gas, so W is **positive**, equal to $p\times$ (decrease in volume).
- **Constant volume**: no movement, so $W = 0$.



On a p–V graph the work done by the gas is the **area under the line**. At constant pressure the area is a rectangle: in the diagram, $1.2 \times 10^5 \times (0.050 - 0.020) = 3600$ J of work is done by the gas, so $W = -3600$ J. This is also how to read the work done in a change that is not at constant pressure.

On a p–V graph:

- **constant volume** is a vertical line;
- **constant pressure** is a horizontal line;
- **constant temperature** (for an ideal gas) is a curve on which pV stays constant: it is steep at small volumes and gets less steep as the volume increases, never touching an axis. Double the volume and the pressure halves.

Special cases of the first law

Process	What is zero	First law gives
Constant volume	$W = 0$	$\Delta U = q$: all the heat goes into internal energy
Constant temperature (ideal gas)	$\Delta U = 0$	$q = -W$: heat supplied equals work done **by** the gas
Very rapid change, or insulated	$q = 0$ (no time for heat to flow)	$\Delta U = W$
Complete cycle	$\Delta U = 0$	net $q =$ − net W

A rapid compression warms a gas. When you pump a tyre quickly, there is no time for thermal energy to leave the air in the pump, so $q = 0$. Work is done **on** the air to compress it, so $W > 0$ and $\Delta U = W > 0$. The internal energy increases, so the temperature rises. A rapid **expansion** does the opposite: the gas does work, $W < 0$, so U falls and the gas cools.

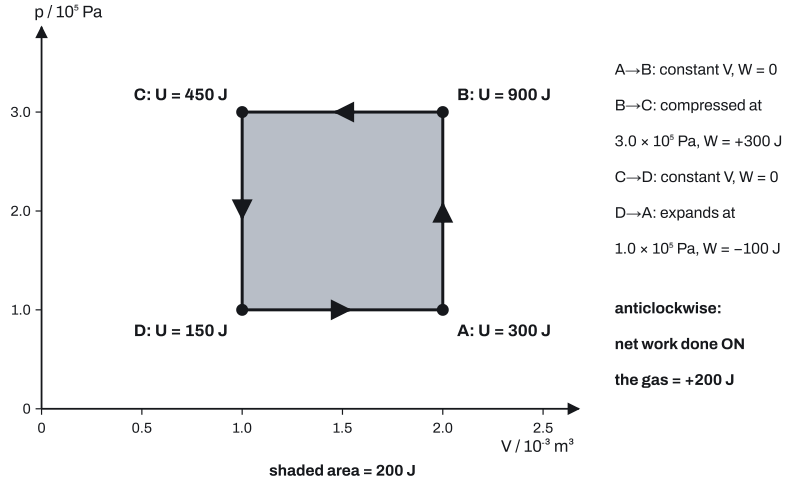
Heating at constant pressure needs more energy than at constant volume. To raise the temperature of an ideal gas by the same amount, ΔU is the same either way (it depends only on the temperature). At constant volume $W = 0$, so $q = \Delta U$. At constant pressure the gas expands and does work, so W is negative and $q =$

$\Delta U - W$ is **larger**. So a gas has a **larger specific heat capacity at constant pressure** than at constant volume.

Boiling at atmospheric pressure. When water boils, the vapour takes up far more space than the liquid, so the water does work pushing back the atmosphere: $W = -p\Delta V$. The latent heat supplied, $q = mL$, is therefore more than the increase in internal energy: $\Delta U = mL - p\Delta V$.

Cycles

A gas taken round a closed loop on a p–V graph returns to its starting state, so over the cycle $\Delta U = 0$, and the net heat supplied equals the net work done **by** the gas.



- The net work is the **area enclosed** by the loop.
- If the loop goes **clockwise**, the gas expands at high pressure and is compressed at low pressure, so it does more work than is done on it: net work done **by** the gas, net heat **supplied**. This is how an engine works.
- If the loop goes **anticlockwise** (as in the diagram), more work is done **on** the gas than it does, so there is a net **output** of heat.

For the diagram, the work done on the gas is $+300$ J from B to C (compressed at 3.0×10^5 Pa through 1.0×10^{-3} m³) and -100 J from D to A, so the net work done on the gas is $+200$ J and the net heat supplied is -200 J: 200 J of heat leaves the gas in each cycle. Worked example 3 fills in the whole table.

Key facts and formulas

"Given" means it is on the data sheet in the exam. "Learn it" means you must remember it. None of this topic's formulas are on the data sheet; the constants R, k and N_A are.

Formula	Status
Internal energy = sum of the random kinetic and potential energies of the molecules	Learn it

Formula	Status
First law: $\Delta U = q + W$ (ΔU increase in internal energy, q thermal energy supplied to the system, W work done on the system)	Learn it
Work done by a gas at constant pressure: $p\Delta V$; work done on the gas $W = -p\Delta V$	Learn it
Work done = area under the p–V graph; net work in a cycle = area enclosed	Learn it
Ideal gas: $U = \frac{3}{2}NkT = \frac{3}{2}nRT = \frac{3}{2}pV$, so $\Delta U = \frac{3}{2}Nk\,\Delta T$	Learn it
Equation of state $pV = nRT = NkT$; fixed mass $\frac{p_1V_1}{T_1} = \frac{p_2V_2}{T_2}$	Learn it
Heating: $q = mc\Delta\theta$; change of state: $q = mL$	Learn it
Complete cycle or constant temperature (ideal gas): $\Delta U = 0$	Learn it

How to do it

In the Paper 4 questions we have tagged for this topic (16 different questions, 2021–2025; 6 more appear on two papers), the types came up this often. 15 of the 16 questions mix two or more types, so the counts add up to more than 16.

Question type	Questions
Explaining a change with the first law	8
Internal energy and molecular energies	7
Energy tables and cycles	6
Work done $p\Delta V$ and its sign	5
Internal energy of an ideal gas from $U = \frac{3}{2}pV$	5
Heating calculations: q, specific heat capacity, latent heat	4
Stating the first law	4
Sketching changes on a graph	3

1. Explaining a change with the first law (8 questions)

1. **Decide each term** of $\Delta U = q + W$ from the situation, with its sign and a reason:
   - q: is thermal energy supplied (positive), removed (negative) or is there no time or no temperature difference for it to flow (zero)?

- W: is the volume decreasing (work done **on** it, positive), increasing (work done **by** it, negative) or constant (zero)?

2. **Combine them** to get the sign of ΔU: say "the internal energy increases" (or decreases).
3. **Link to what is asked**: temperature rises if the kinetic energy of the molecules increases; for an ideal gas, $U \propto T$.

Variations you will meet:

- **A rapid compression** (a pump, a piston pushed in quickly): $q = 0$ because there is no time for heat transfer; work is done on the gas; so U increases and the temperature rises.
- **A sudden expansion with no heat transferred**: $q = 0$, so $\Delta U = W$; the gas does work, so W is negative, U falls and the gas cools.
- **A spring or wire stretched slowly at constant temperature**: no heat transferred, work is done on it, so its internal energy increases (as potential energy of the molecules).
- **Water evaporating on a hot day**: heat is supplied to the water and the water does work pushing back the atmosphere as it becomes vapour, but more heat goes in than work comes out, so U increases.
- **Compare two specific heat capacities** (constant pressure against constant volume): same temperature rise, so same ΔU; at constant volume no work is done; so **less** heat is needed and the specific heat capacity at constant volume is **smaller**. Say all three steps.
- **How a specific latent heat changes at a higher pressure**: name all three terms in one line of reasoning. At the higher pressure, more work is done **by** the water pushing back the surroundings as it expands into vapour, so W (work done on it) is more negative; the increase in internal energy ΔU for the change of state is about the same; so $q = \Delta U - W$ is larger: more thermal energy is needed per kilogram, and L **increases**. Say "increases" or "decreases", never just "changes".
- **A change on a p–V graph** (for example "explain how the first law applies to Z to X"): say whether the volume increases (work done **by** the gas) and whether pV, and so T and U, increases or decreases. Say "increases" or "decreases", never just "changes".

2. Internal energy and molecular energies (7 questions)

1. **Definition**: the sum of the (random) kinetic and potential energies of the molecules. Both halves earn marks: "sum of kinetic and potential energy" and "random motion of the molecules".
2. **For a process**, deal with kinetic and potential energy separately:
   - temperature changes → kinetic energy changes; temperature constant → kinetic energy constant;
   - separation changes → potential energy changes; separation constant → potential energy constant.
3. **Conclude** about the internal energy.

Variations you will meet:

- **Internal energy of an ideal gas**: the total kinetic energy of the random motion of the molecules, **because** there are no intermolecular forces, so the potential energy is zero.
- **Why $U \propto T$ for an ideal gas**: potential energy is zero, so U is the total kinetic energy; the mean kinetic energy of a molecule is $\frac{3}{2}kT$, proportional to T.

- **Heating a gas at constant volume**: same separation so no change in potential energy; temperature rises so kinetic energy rises; so U increases.
- **Boiling or melting at constant temperature**: the separation of the molecules increases so potential energy increases; the temperature is constant so kinetic energy is unchanged; that is why the temperature does not rise although energy is supplied.
- **Stretching a wire at constant temperature**: kinetic energy unchanged, separation increases so potential energy increases, so U increases.

3. Energy tables and cycles (6 questions)

These give a table with columns "work done on gas", "thermal energy supplied to gas" and "increase in internal energy", one row per change.

1. **Fill in what is given first**, with signs. Energy removed or work done **by** the gas is **negative**.
2. **Fill in the zeros**: constant volume → work done $= 0$; constant temperature → increase in internal energy $= 0$.
3. **Use $\Delta U = q + W$ across each row** to get the third entry.
4. **Round the cycle**, the internal energy column adds up to **zero**. Use this to get a missing ΔU.
5. **Net heat over a cycle** $= -$(net work done on the gas) $=$ net work done by the gas. Say why: $\Delta U = 0$ for the cycle.

Variations you will meet:

- **Tables in letters**: if a gas has internal energy U at temperature T, then because $U \propto T$ it has $2U$ at $2T$. So a change from T to $2T$ has $\Delta U = +U$; if work W is done on the gas during it, $q = \Delta U - W = U - W$. Cooling it next at constant volume from $2T$ to $1.5T$: $W = 0$, $\Delta U = 1.5U - 2U = -0.5U$, so $q = -0.5U$.
- **Given work in symbols from a graph**: $W = p\Delta V$ with the letters on the axes, for example a compression at pressure $3Y$ from volume $5X$ to $3X$ has $W = +3Y \times 2X = +6XY$. Use only letters defined in the question, and use p only for a pressure, never as an energy.
- **Signs from a p–V graph** (ticks in a table): U increases where pV increases; work done on the gas is positive where V decreases, zero where V is constant.
- **Explaining the energy transfers in a whole cycle**: $\Delta U = 0$; the gas does more work in the expansion (at the higher pressure) than is done on it in the compression; so there must be a net input of thermal energy equal to that net work. Heat and work are **not** zero over a cycle; only ΔU is.
- **ΔU for one complete cycle**: always 0.

4. Work done $p\Delta V$ and its sign (5 questions)

1. **Find ΔV** = final volume – initial volume, in m³ (1 cm³ $= 10^{-6}$ m³). Show the subtraction.
2. **Multiply by the (constant) pressure** in Pa: $p\Delta V$ is the work done **by** the gas.
3. **State the sign**: expanding → work done on the gas is negative; compressed → positive. Give the reason ("the gas expands, so it does work on the surroundings").

Variations you will meet:

- **"Show that the magnitude of the work done is …"**: write $W = p\Delta V$, substitute including the subtraction of volumes, and give the answer.
- **Volume change from temperature at constant pressure**: first use $\frac{V_1}{T_1} = \frac{V_2}{T_2}$ (in kelvin) to find the new volume, then $p\Delta V$.
- **Vapour from a liquid**: the liquid's volume is negligible, so ΔV is the vapour's volume from $pV = nRT$ with $n = \frac{\text{mass}}{\text{molar mass}}$.
- **From a graph**: the area under a horizontal line. A constant-volume stage has $W = 0$.

5. Internal energy of an ideal gas from $U = \frac{3}{2}pV$ (5 questions)

1. **Say why**: the potential energy is zero, so U is the total kinetic energy $= N \times \frac{3}{2}kT$.
2. **Use** $pV = NkT$ to write it as $U = \frac{3}{2}pV$ (no temperature needed) or $U = \frac{3}{2}nRT$.
3. **For a change**: $\Delta U = \frac{3}{2}Nk\,\Delta T$ or $\frac{3}{2}\Delta(pV)$, with the sign.

Variations you will meet:

- **Deduce the expression from a graph** of U against time and a p–V graph: compare U with pV at each state, find $U = \frac{3}{2}pV$, then use $pV = NkT$ (with k the Boltzmann constant) to get $U = \frac{3}{2}NkT$.
- **A gas that cools** has a **negative** increase in internal energy, for example -1400 J. Multiply by N: the change for one molecule is not the answer.
- **Total kinetic energy change** after a rapid compression: find N from $pV = NkT$ in the first state, then $\frac{3}{2}kN\Delta T$ (a temperature **change**: no $+273$).

6. Heating calculations: q, specific heat capacity, latent heat (4 questions)

1. **Find the missing energy with the first law**: $q = \Delta U - W$. Watch the sign of W: for an expanding gas, $q = \Delta U + p\Delta V$.
2. **Specific heat capacity** of the gas for that process: $c = \frac{q}{m\Delta\theta}$, with m in kg and the unit J kg^{-1} K^{-1}.
3. **Latent heat**: $q = mL$; then $\Delta U = mL - p\Delta V$ if the substance expands against the atmosphere.

Variations you will meet:

- **Energy removed to cool a gas and condense it**: $mc\Delta\theta$ for the cooling (with $\Delta\theta$ from the start to the boiling point) **plus** mL for the condensing. Keep units consistent (kJ with kJ kg^{-1}).
- **Compare** with the constant-volume case (type 1): same ΔU, no work, less heat, smaller c.

7. Stating the first law (4 questions)

1. **Words**: the increase in internal energy of a system equals the thermal energy supplied to the system by heating plus the work done on the system.

2. **Symbols**: $\Delta U = q + W$, with each symbol named **with its direction** ("increase in", "supplied to", "done on").

Variations you will meet:

- **"State two ways the internal energy may be changed"**: by work done on (or by) the system; by thermal energy transferred to (or from) the system.
- **Meaning of ΔU, q, W**: say "increase in internal energy of the system", "thermal energy supplied **to** the system", "work done **on** the system". Giving no direction, or both directions, loses the mark.

8. Sketching changes on a graph (3 questions)

1. **Plot each state** you know, then join them with the right kind of line: constant V vertical, constant p horizontal, constant T a curve with pV constant.
2. **Find missing states first** with $pV \propto T$: at constant volume, $p \propto T$ (double the pressure, double the temperature); at constant temperature, pV is constant (halve the pressure, double the volume).
3. **Label** the lines or states as asked.

Variations you will meet:

- **Constant-temperature curve**: its gradient is negative and gets **less steep** as V increases. Draw a curve, not a straight line, and end it exactly at the two states.
- **U against V at constant pressure**: $U = \frac{3}{2}pV$, a straight line through the origin.

Worked examples

The bold codes show where marks are earned, as in Cambridge mark schemes. The number is how many marks.

- **A1** accuracy mark: for a correct answer; it needs the M mark before it.
- **B1** mark for a correct result on its own, with no method needed.

Example 1

An ideal gas at a temperature of 300 K expands at a constant pressure of 1.2×10^5 Pa from a volume of 0.020 m³ to 0.050 m³ as it is heated.

(a) Calculate the work done on the gas. **[2]**

(b) Calculate the final temperature of the gas. **[1]**

(c) Show that the internal energy of the gas increases by 5400 J. **[2]**

(d) Use the first law of thermodynamics to determine the thermal energy supplied to the gas. **[2]**

Solution

(a) Work done by the gas $= p\Delta V = 1.2 \times 10^5 \times (0.050 - 0.020) = 3600$ J **C1** The gas expands, so the work done **on** it is $W = -3600$ J **A1**

(b) At constant pressure $V \propto T$: $T = 300 \times \dfrac{0.050}{0.020} = 750$ K **A1**

(c) For an ideal gas $U = \frac{3}{2}pV$, so at constant pressure $\Delta U = \frac{3}{2}p\Delta V$ **C1** $\Delta U = 1.5 \times 1.2 \times 10^5 \times 0.030 = 5400$ J **A1**

(d) $\Delta U = q + W$, so $5400 = q + (-3600)$ **C1** $q = 9000$ J **A1**

Example 2

(a) State what is meant by the internal energy of a system. **[2]**

(b) A block of ice at $0\,^\circ$C melts to water at $0\,^\circ$C. With reference to molecular energies, explain how the internal energy changes. **[3]**

(c) A gas in a syringe is allowed to expand very rapidly, pushing the plunger out. Use the first law of thermodynamics to explain why the temperature of the gas falls. **[3]**

Solution

(a) The sum of the kinetic and potential energies of the molecules **B1**, spread at random among them (the energy of their random motion) **B1**.

(b) The temperature is constant, so the mean kinetic energy of the molecules does not change **B1**. The molecules move apart from their fixed positions, so their potential energy increases **B1**. Kinetic energy is unchanged and potential energy increases, so the internal energy increases **B1**.

(c) The expansion is so rapid that there is no time for thermal energy to be transferred, so $q = 0$ **B1**. The gas expands, so it does work on the plunger: the work done on the gas is negative **B1**. So $\Delta U = W$ is negative: the internal energy, and so the kinetic energy of the molecules and the temperature, decreases **B1**.

Example 3

A fixed mass of ideal gas is taken round the cycle ABCDA shown in the third diagram in "Understand it". At A the gas has volume 2.0×10^{-3} m³, pressure 1.0×10^5 Pa and temperature 300 K. B is at 2.0×10^{-3} m³ and 3.0×10^5 Pa, C at 1.0×10^{-3} m³ and 3.0×10^5 Pa, and D at 1.0×10^{-3} m³ and 1.0×10^5 Pa.

(a) Determine the temperature at B. **[1]**

(b) Show that the internal energy of the gas at A is 300 J, and find it at B, C and D. **[2]**

(c) Complete a table giving, for each change, the work done on the gas, the thermal energy supplied to the gas and the increase in internal energy. **[5]**

(d) Determine the net thermal energy supplied to the gas in one cycle. Explain your reasoning. **[2]**

Solution

(a) Constant volume, so $p \propto T$: $T_B = 300 \times 3 = 900$ K **A1**

(b) $U = \frac{3}{2}pV = 1.5 \times 1.0 \times 10^5 \times 2.0 \times 10^{-3} = 300$ J at A **C1** At B: 900 J; at C: 450 J; at D: 150 J **A1**

(c) Work: AB and CD are at constant volume, so $W = 0$ **B1**. BC is a compression at 3.0×10^5 Pa through 1.0×10^{-3} m³: $W = +300$ J. DA is an expansion at 1.0×10^5 Pa: $W = -100$ J **B1** Increase in internal energy from (b): AB $+600$, BC -450, CD -300, DA $+150$ J (these add up to 0) **B1** Thermal energy from $q = \Delta U - W$ in each row **B1 B1**:

Change	Work done on gas / J	Thermal energy supplied / J	Increase in internal energy / J
A to B	0	$+600$	$+600$
B to C	$+300$	-750	-450
C to D	0	-300	-300
D to A	-100	$+250$	$+150$

(d) Over the cycle the gas returns to A, so $\Delta U = 0$ **B1**; net work done on the gas $= 300 - 100 = +200$ J, so net thermal energy supplied $= -200$ J (a net 200 J of heat leaves the gas) **A1**. Check: $600 - 750 - 300 + 250 = -200$ J.

Example 4

0.50 mol of helium, of mass 2.0 g, may be treated as an ideal gas. It is at a pressure of 1.0×10^5 Pa and a temperature of 300 K.

(a) Calculate the volume of the gas. **[2]**

(b) The gas is heated at constant pressure to 360 K. Calculate the work done on the gas. **[2]**

(c) Calculate the increase in internal energy of the gas. **[2]**

(d) Determine the specific heat capacity of helium for this process. **[3]**

(e) The same gas is now heated from 300 K to 360 K at constant volume. Use the first law of thermodynamics to explain whether its specific heat capacity is greater or smaller, and calculate it. **[3]**

Solution

(a) $pV = nRT$, so $V = \dfrac{0.50 \times 8.31 \times 300}{1.0 \times 10^5}$ **C1** $V = 1.25 \times 10^{-2}$ m³ **A1**

(b) $V \propto T$, so $\Delta V = 1.2465 \times 10^{-2} \times \dfrac{60}{300} = 2.493 \times 10^{-3}$ m³ **C1** Work done by the gas $= p\Delta V = 249$ J; the gas expands, so $W = -249$ J **A1**

(c) $\Delta U = \frac{3}{2}nR\,\Delta T = 1.5 \times 0.50 \times 8.31 \times 60$ **C1** $\Delta U = 374$ J **A1**

(d) $q = \Delta U - W = 373.95 + 249.3 = 623$ J **C1** $c = \dfrac{q}{m\Delta\theta} = \dfrac{623.25}{0.0020 \times 60}$ **C1** $c = 5190$ J kg^{-1} K^{-1} **A1**

(e) Same temperature rise, so the same increase in internal energy, 374 J **B1**. At constant volume no work is done, so $q = \Delta U = 374$ J: less thermal energy is needed **B1**. So the specific heat capacity is smaller: $c = \dfrac{373.95}{0.0020 \times 60} = 3120$ J kg^{-1} K^{-1} **B1**.

Common mistakes

- **Stating the first law without directions.** The November 2021 report says very few stated the law in words, and some wrote "internal energy" instead of "**increase** in internal energy". The June 2023 report says only a minority gave the direction of all three quantities consistently; the March 2022 report saw candidates giving both directions or none. Say "increase in", "supplied **to**", "done **on**".
- **Mixing up the definition of internal energy and the first law.** The June 2021 report describes many "first laws" written in terms of internal energy itself rather than its change. Internal energy is the kinetic plus potential energy of the molecules; the first law is how it can be **changed**.
- **Treating heat and work as energy the system has.** The June 2023 report stresses that thermal energy and work are transfers, not stores, so they cannot "change". The March 2023, November 2022 and March 2025 reports all describe answers that confused internal energy with thermal energy.
- **"ΔU increases".** The March 2023 report says many treated U and ΔU as the same thing. Write "U increases" or "ΔU is positive".
- **Wrong sign for work done on an expanding gas.** The March 2025 report says only the strongest candidates realised that the work done on an expanding gas is negative. The June 2022 report saw "the work done by the water is negative", which is wrong: work done **by** it is positive, work done **on** it is negative. Say which one you mean.
- **Zero in the wrong column.** The June 2022 report describes tables with zero heat and a 31.6 kJ increase in internal energy for a change at constant temperature. At constant temperature an ideal gas has $\Delta U = 0$, so the heat equals the work done by the gas.
- **Thinking heat and work are also zero over a cycle.** The June 2023 report says candidates who spotted $\Delta U = 0$ for a cycle often said the heat and work were zero too. Only ΔU is zero; the net heat supplied equals the net work done by the gas.
- **Not starting a table from the given entries.** The June 2023 report says many didn't know where to begin: fill in what is given, then the zeros, then use $\Delta U = q + W$ row by row and the zero total of ΔU. It also saw letters that were not defined, or p used as an energy.
- **One molecule instead of the whole gas.** The June 2021 report says many found the kinetic energy change of a single molecule, and few noticed the internal energy was **decreasing**, so its increase was negative.
- **Not reading U from a p–V graph.** The June 2023 report says many did not see that $U \propto pV$ for an ideal gas, so its direction of change can be read straight off the graph.
- **Specific heat capacity comparisons.** The November 2022 report says the reasoning was poor, again from confusing thermal energy with internal energy. Use three steps: same ΔU; different work; so different heat needed.

- **A straight line for a constant-temperature change.** The June 2023 report says the line joining two states at the same temperature was often drawn straight; it must be a curve with pV constant.
- **Using the first law when kinetic theory is asked for.** The June 2021 report says many explained cooling on expansion with the first law when the question asked for kinetic theory (molecules rebounding more slowly from a retreating piston). Answer with the idea the question names.

How to get the marks

- **Definitions** (one mark per idea):
  - Internal energy: sum of the kinetic and potential energies of the molecules; energy of their **random** motion.
  - Internal energy of an ideal gas: total kinetic energy of the random motion of the molecules, **because** the potential energy is zero.
  - First law: increase in internal energy = thermal energy supplied to the system + work done on the system (words, or symbols each defined with its direction).
- **"Explain" with the first law**: a mark for each term with a reason (q zero because there is no time for heat to flow; work done on it because it is compressed), then a mark for the conclusion about U and the temperature. Say "increases" or "decreases", not "changes".
- **Molecular energy explanations**: one mark for kinetic energy (with the temperature reason), one for potential energy (with the separation reason), one for the conclusion.
- **Calculations (C1 then A1)**: write $W = p\Delta V$ or $\Delta U = q + W$, substitute with signs, then give the answer with its unit (J, or J kg^{-1} K^{-1}).
- **"Show that"**: show the subtraction of volumes and every substitution; end at the given value.
- **Tables**: each correct entry or column earns marks, and follow-through is given if each row obeys $\Delta U = q + W$ and the ΔU column adds to zero round a cycle. Include the $+$ or $-$ sign on every non-zero entry.
- **"Explain your reasoning"** (net heat in a cycle): state $\Delta U = 0$ for the cycle, and that the work in the stages at constant volume is zero, before the number.
- **Sketches**: a mark for each correct line: the right shape (vertical, horizontal, or a curve with decreasing gradient) between the right two points, labelled as asked.
- **Significant figures**: match the data (usually 2 or 3 significant figures). Keep unrounded values between parts, as in Worked example 4.

Quick check

1. A gas is compressed at a constant pressure of 2.0×10^5 Pa from 3.0×10^{-3} m³ to 1.8×10^{-3} m³. During the compression, 500 J of thermal energy leaves the gas. Find the work done on the gas and the increase in its internal energy.
2. An ideal gas expands slowly at constant temperature, doing 150 J of work on its surroundings. Find the thermal energy supplied to the gas.

3. 0.10 kg of water boils at $100\,^{\circ}$C at a pressure of 1.0×10^{5} Pa. The specific latent heat of vaporisation is 2.26×10^{6} J kg^{-1}, and 1 mol of water has mass 18 g. Treat the vapour as an ideal gas and the liquid's volume as negligible. Find the thermal energy supplied, the work done on the water, and the increase in its internal energy.
4. A metal wire is stretched slowly at constant temperature, within its elastic limit. With reference to molecular energies, explain what happens to its internal energy.
5. A gas is taken round a cycle in which it does 80 J more work than is done on it. Explain why the net thermal energy supplied to the gas in one cycle is 80 J.

Answers

1. Work done on the gas $= p\times$ decrease in volume $= 2.0 \times 10^5 \times 1.2 \times 10^{-3} = +240$ J (it is compressed, so positive). $\Delta U = q + W = -500 + 240 = -260$ J: the internal energy decreases by 260 J.
2. Constant temperature, so $\Delta U = 0$. The gas does work, so $W = -150$ J. $q = \Delta U - W = 0 + 150 = +150$ J.
3. $q = mL = 0.10 \times 2.26 \times 10^6 = 2.26 \times 10^5$ J. $n = \dfrac{100}{18} = 5.56$ mol, so $V = \dfrac{nRT}{p} = \dfrac{5.556 \times 8.31 \times 373}{1.0 \times 10^5} = 0.172$ m³. The water does work $p\Delta V = 1.0 \times 10^5 \times 0.172 = 1.72 \times 10^4$ J pushing back the atmosphere, so $W = -1.72 \times 10^4$ J. $\Delta U = q + W = 2.26 \times 10^5 - 1.72 \times 10^4 = 2.09 \times 10^5$ J.
4. The temperature is constant, so the kinetic energy of the molecules does not change. The molecules are pulled further apart, so their potential energy increases. So the internal energy increases. (First law: no heat transferred, work done on the wire, so $\Delta U = W > 0$.)
5. The gas returns to its starting state, so $\Delta U = 0$ for the cycle. The net work done on the gas is -80 J, so $0 = q + (-80)$, and $q = +80$ J: a net 80 J of thermal energy is supplied.

Past questions to try

- **9702/43/M/J/23 Q4**: a three-stage change sketched on a p–V graph, then an energy table in symbols.
- **9702/43/M/J/22 Q3**: showing the work done at constant pressure, then an energy table in kJ for a three-stage cycle.
- **9702/42/M/J/22 Q3**: latent heat, the work done by water as it boils, and the first law at a higher pressure.
- **9702/43/O/N/23 Q2**: work done on an expanding gas, its sign, the heat supplied and a specific heat capacity comparison.