

Thermodynamics

A2 Physics ■ Topic 16

A-Level Physics



What we will cover

- 1 Internal energy 16.1
- 2 The first law $\Delta U = q + W$ 16.2
- 3 Recap



1

Internal energy

Sheet 16.1 – what you need to know

You must be able to:

- state what **internal energy** is
- explain why an ideal gas has no potential energy, so $U \propto T$
- find U from N and the mean kinetic energy, or from $U = \frac{3}{2}pV$
- say what happens to U when a substance changes state

What the exam asks

“State what is meant by the internal energy of a system.” [2]

“Use one of the basic assumptions of the kinetic theory to explain what can be deduced about the potential energy in an ideal gas.” [2]

“Sketch the variation with V of the internal energy U of the gas.” [2]

Practise on exercise sheet 16.1, then try the real exam questions on past-paper sheet 16.1.

Internal energy

Internal energy 内能 U is the sum of the random kinetic energy of the molecules and the potential energy stored in the intermolecular forces. It is a **state function**: it depends only on the present state, not on how the system got there.

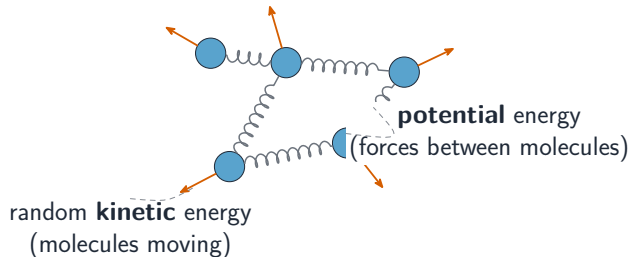
random KE

motion of molecules

molecular PE

bonds / attractions

The two contributions



The **random kinetic energies** 无规则动能 of the molecules – **translational** 平动 as they fly about, and for anything but a **monatomic** 单原子 gas also **rotational** 转动 and **vibrational** 振动 – plus the potential energy in the intermolecular forces. Not the body's own motion, and not its position.

Two things U is not

U depends only on the *state* – temperature, pressure, volume and **amount of substance** 物质的量 – never on the path taken to reach it.

Not path-dependent

U depends only on the *state*
– T , p , V , amount – never
on how the gas got there.

Not bulk motion

a moving cylinder of gas has
kinetic energy, but U counts only
the *random* molecular motion.

Throw a sealed flask across the room and its U does **not** change – the molecules are no more agitated than before.

Temperature and internal energy

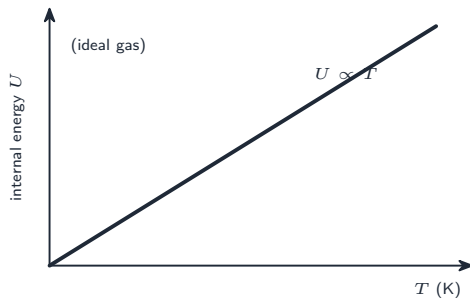
Each ideal-gas molecule carries $\frac{3}{2}kT$ on average (Topic 15), and there is no molecular PE, so

$$U = \frac{3}{2}NkT = \frac{3}{2}nRT$$

$U \propto T$: double the thermodynamic temperature, double the internal energy.

This is **exact only for an ideal gas**. In a **phase change** 相变, U climbs while T stands still – the energy goes into **potential**, breaking bonds.

Temperature and internal energy, in detail



For an ideal gas $U \propto T$ – a straight line through the origin. A **phase change** breaks it: U climbs while T stands still.

Internal energy: the answers that score

“State what is meant by the internal energy of a system.” [2]

The **sum** of the **random** kinetic and potential energies of its molecules.

“Why is there no potential energy in an ideal gas?” [2]

No **intermolecular forces** (except in collisions), so the potential energy is **zero**: U is all kinetic.

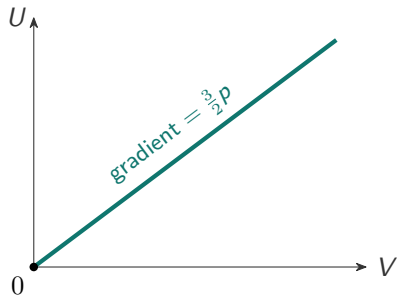
A gas at 1.0×10^5 Pa fills 2.0×10^{-3} m³ at 300 K. Find U .

$$N = \frac{pV}{kT} = \frac{1.0 \times 10^5 \times 2.0 \times 10^{-3}}{1.38 \times 10^{-23} \times 300} \\ = 4.83 \times 10^{22}$$

$$\langle E_K \rangle = \frac{3}{2} kT = \frac{3}{2} \times 1.38 \times 10^{-23} \times 300 \\ = 6.21 \times 10^{-21} \text{ J}$$

$$U = N \langle E_K \rangle = 300 \text{ J}$$

Internal energy against volume at constant pressure



For an ideal gas $U = \frac{3}{2}NkT$, and $NkT = pV$:

$$U = \frac{3}{2}pV$$

- at constant p : $U \propto V$ – a straight line through the origin
- expanding at constant p raises T , so it raises U

The same $U = \frac{3}{2}pV$ gives U at any point of a p – V diagram: no temperature needed.

Heating without a temperature rise

A phase change

Supply heat at constant pressure while water boils: the temperature stays constant, yet the internal energy still rises.

Where it went

The **latent heat** 潜热 separates the molecules – it increases the potential-energy part of U , not the kinetic part.

And the gas expands

The vapour occupies far more volume, so the system also does expansion work $p\Delta V$ on the atmosphere.

$\Delta U = q + W$ still holds, with $W = -p\Delta V$ (the vapour does work on the atmosphere). A constant temperature means constant *kinetic* energy, never constant internal energy.

2

The first law

Sheet 16.2 – what you need to know

You must be able to:

- find the work done from $p\Delta V$ or the area under a p - V graph
- use $\Delta U = q + W$ with W the work done on the gas
- apply it to compression, expansion, heating and cycles
- explain why heating at constant pressure needs more energy

What the exam asks

“Use the first law of thermodynamics to explain why a bicycle pump gets hot when it is used to pump up a tyre quickly.”

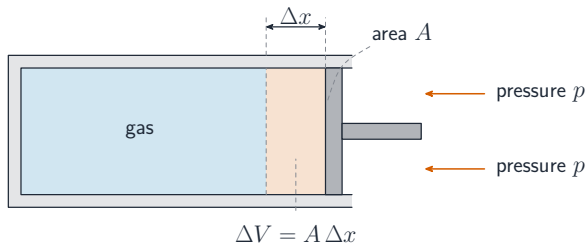
[3]

“Explain whether the work done on the gas is positive or negative.” **[2]**

“Show that the helium gas behaves as an ideal gas.” **[2]**

Practise on exercise sheet 16.2, then try the real exam questions on past-paper sheet 16.2.

Work done by a gas

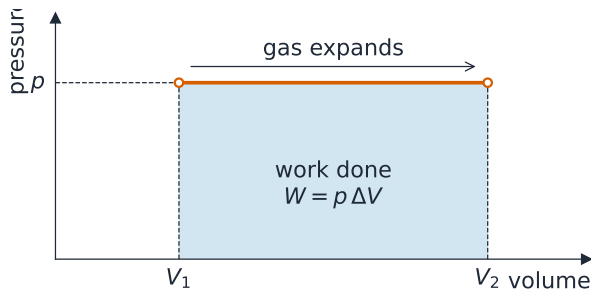


A gas pushing a **piston** 活塞 of area A out by Δx sweeps $\Delta V = A\Delta x$. At constant pressure the work is

$$W = p \Delta V$$

It follows straight from $W = Fx$: the force is pA and the distance Δx , so $W = pA\Delta x = p\Delta V$.

Work is the area under a p - V graph



$$\text{work} = \text{area under } p\text{-}V \cdot W = \int p dV$$

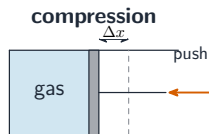
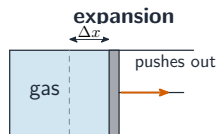
At constant pressure the area under the line is a rectangle, $p \Delta V$.

Gas at $1.0 \times 10^5 \text{ Pa}$ expands from 2.0×10^{-3} to $5.0 \times 10^{-3} \text{ m}^3$. Find the work done *by* it.

$$W = p \Delta V = (1.0 \times 10^5)(3.0 \times 10^{-3}) = 300 \text{ J}$$

On a pressure–volume graph, the work done at constant pressure is the area under the line: $W = p \Delta V$

The sign convention

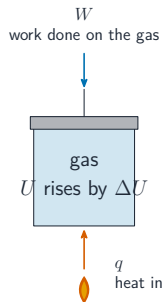


This syllabus writes $\Delta U = q + W$ with W the work done **on** the gas:

- **compressed**: $\Delta V < 0$, $W > 0$ – gas gains
- **expands**: $\Delta V > 0$, $W < 0$ – gas loses
- $\Delta V = 0$: no work at all

Read the question's preposition.
Work **on** the gas and work **by** the gas are the same size and **opposite** sign – most lost marks in this topic live here.

The first law of thermodynamics



$$\Delta U = q + W$$

both are positive when
energy goes *into* the gas

Conservation of energy 能量守恒,
written for heat and work crossing a
boundary:

$$\Delta U = q + W$$

- q : energy added **by heating**
- W : work done **on** the gas

Both terms are **positive** when energy **goes in**. That single rule fixes every sign in the topic.

Reading the equation

$$\Delta U = q + W$$

ΔU is fixed by the change of state; for an ideal gas, by the change in T

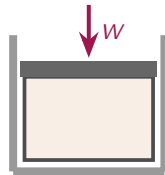
rigid: constant volume



all heat, no work

$$\Delta U = q$$

insulated walls



all work, no heat

$$\Delta U = W$$

The first law at work



A turbine takes in heat, does work, and rejects the rest – $\Delta U = q + w$ with every term non-zero.

Worked example – expanding while heated

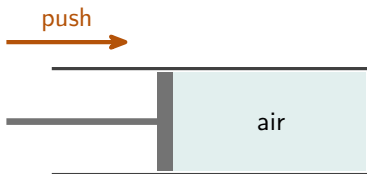
A gas absorbs 500 J of heat while doing 200 J of work on its surroundings. Find ΔU .

The gas **does** the work, so the work done *on* it is negative: $W = -200 \text{ J}$.

$$\Delta U = q + W = 500 + (-200) = 300 \text{ J}$$

Translate the words **before** substituting: “does work on its surroundings” is the phrase that flips the sign.

The bicycle pump



$W > 0$: work done on the air

$q \approx 0$: too quick to lose heat

$\Delta U = q + W > 0$: T rises

“Use the first law of thermodynamics to explain why a bicycle pump gets hot when it is used to pump up a tyre quickly.” [3]

The air is compressed: work is done on it, so W is positive.

It happens quickly, so there is no time for heat to escape: $q \approx 0$.

So $\Delta U = q + W > 0$: the internal energy, and the temperature, rise.

Same ΔU , different routes

All heat, no work

$\Delta U = q$ – heating a
sealed rigid container

All work, no heat

$\Delta U = W$ – compress-
ing an insulated cylinder

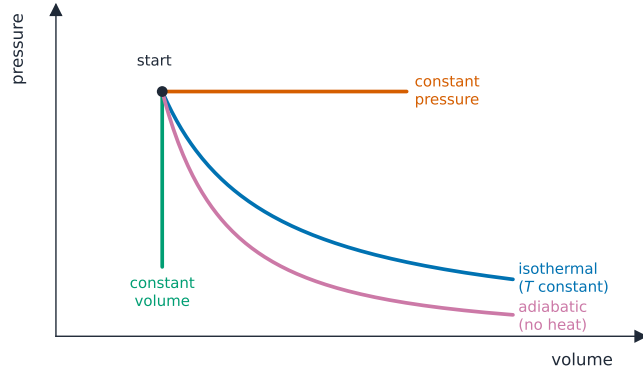
ΔU is fixed by the **change of state** (for an ideal gas, by ΔT alone). The **mix** of q and W that delivers it is up to the process.

The four standard processes

Process	fixed	ΔU	then
Isothermal 等温	T	0	$q = -W$
Constant V	V	$\frac{3}{2}nR\Delta T$	$q = \Delta U$
Constant p	p	$\frac{3}{2}nR\Delta T$	$W = -p\Delta V$
Adiabatic 绝热	$q = 0$	W	–

Read every row off $\Delta U = q + W$. Isothermal: heat in comes **straight back out** as work.
 Adiabatic: the gas warms **only** because work is done on it.

The four standard processes, in detail



isothermal, isobaric, isometric on p - V plane

Isothermal: heat in comes straight back out as work.

Worked example – a two-step process

Gas at T with internal energy U : (1) compress to $3T$, work W done on it; (2) cool at constant volume to $2T$.

Since $U = \frac{3}{2}nRT$, U tracks T :

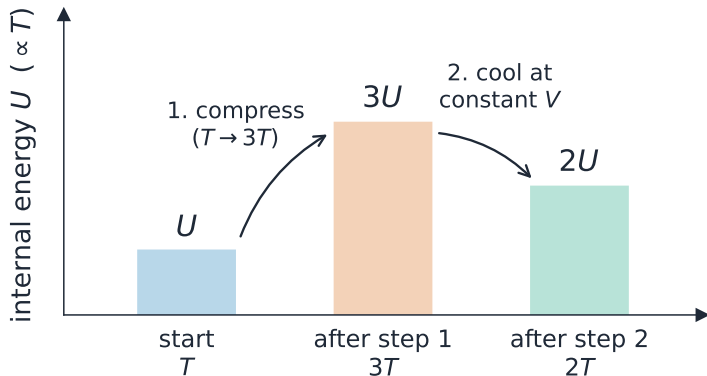
1. $U \rightarrow 3U$, so $\Delta U_1 = 2U$, $W_1 = +W$

$$\Rightarrow q_1 = 2U - W$$

2. $\Delta U_2 = -U$, $W_2 = 0 \Rightarrow q_2 = -U$ (out)

Check: total $\Delta U = 2U - U = U$, i.e. $T \rightarrow 2T$. Consistent.

Worked example – a two-step process, in detail



ideal gas: $U = \frac{3}{2}nRT$ · internal energy depends only on T

Internal energy tracks temperature: $U \rightarrow 3U$ on compressing to $3T$, then $3U \rightarrow 2U$ on cooling to $2T$

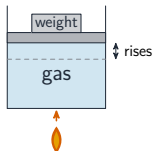
Constant volume vs constant pressure

constant volume



heat q
no work done
all the heat raises U : $q = \Delta U$

constant pressure



heat q
heat raises U and
does work lifting the weight

- constant V : no work, so **all** the heat raises $U - q = \frac{3}{2}nR\Delta T$
- constant p : the heat does **two** jobs – raise U and push the piston out

So heating a gas through the same ΔT costs **more** at constant pressure. Boiling water is the limit case: T holds still, U still climbs, and the vapour does expansion work – and $\Delta U = q + W$ holds throughout.

Worked example – the same ΔU at constant p and at constant V

A gas is heated at $2.0 \times 10^5 \text{ Pa}$. Its internal energy rises by 7600 J and its volume by $1.2 \times 10^{-2} \text{ m}^3$. Find the heat supplied, then the heat needed for the same ΔU at constant volume.

Constant pressure: the gas does work $p\Delta V$, so

$$W = -p\Delta V = -(2.0 \times 10^5 \text{ Pa})(1.2 \times 10^{-2} \text{ m}^3) = -2400 \text{ J}$$

$$q = \Delta U - W = 7600 \text{ J} + 2400 \text{ J} = 1.0 \times 10^4 \text{ J}$$

Constant volume

$$\Delta V = 0 \Rightarrow W = 0$$

$$q = \Delta U = 7600 \text{ J}$$

Less heat: none is spent pushing back the surroundings.

So the specific heat capacity of a gas is **larger at constant pressure** than at constant volume.

Checking that a gas behaves ideally

**Helium goes from $1.0 \times 10^5 \text{ Pa}$, $3.0 \times 10^{-3} \text{ m}^3$, 300 K to $2.4 \times 10^5 \text{ Pa}$, $1.5 \times 10^{-3} \text{ m}^3$, 360 K .
Show that it behaves as an ideal gas.**

A fixed mass of ideal gas keeps pV/T constant:

$$\frac{p_1 V_1}{T_1} = \frac{1.0 \times 10^5 \times 3.0 \times 10^{-3}}{300} = 1.00 \text{ J K}^{-1}$$

$$\frac{p_2 V_2}{T_2} = \frac{2.4 \times 10^5 \times 1.5 \times 10^{-3}}{360} = 1.00 \text{ J K}^{-1}$$

Equal, so the gas is **ideal** over this change.

Then find the rise in the total kinetic energy of its atoms.

$$\Delta U = \frac{3}{2} Nk \Delta T \text{ and } Nk = \frac{pV}{T}:$$

$$\Delta U = \frac{3}{2} \times 1.00 \text{ J K}^{-1} \times 60 \text{ K} = 90 \text{ J}$$

3

Recap

Exam tips

- **First law:** $\Delta U = q + W$ – be careful with the sign convention (W is the work done **on** the gas).
- For an ideal gas, **internal energy depends only on temperature** ($\Delta U \propto \Delta T$).
- Work done by a gas at constant pressure = $p\Delta V$; read the sign from expansion (by) or compression (on).

Topic 16 – key takeaways

1 Internal energy

random KE + molecular PE; a *state function*

2 Ideal gas

$$U = \frac{3}{2}nRT - \text{depends on } T \text{ alone}$$

3 Work

$W = p\Delta V$; on the gas is + when compressed

4 First law

$\Delta U = q + W$ – energy conservation for a gas

Check yourself

- 1 A gas is compressed. Sign of W in $\Delta U = q + W$?
- 2 Isothermal change: what is ΔU ?
- 3 Adiabatic change: what is q ?
- 4 Does a phase change raise U at constant T ?



Answers 1. positive 2. zero 3. zero 4. yes – the PE rises