

Thermodynamics

A2 Physics ■ Topic 16

A-Level Physics



What we will cover

- 1 Internal energy
- 2 Work done on a gas
- 3 The sign convention
- 4 The first law $\Delta U = q + W$
- 5 The standard processes
- 6 Recap





Internal energy

Internal energy 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.

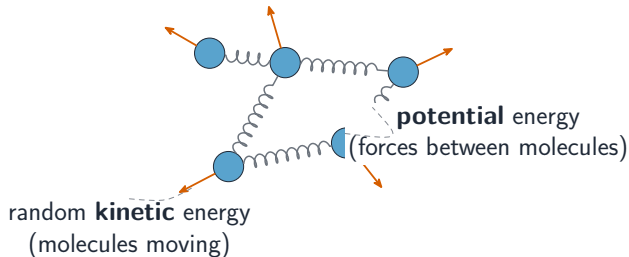
random KE

motion of molecules

molecular PE

bonds / attractions

The two contributions



The **random kinetic energies** of the molecules – **translational** as they fly about, and for anything but a **monatomic** gas also **rotational** and **vibrational** – 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** – 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

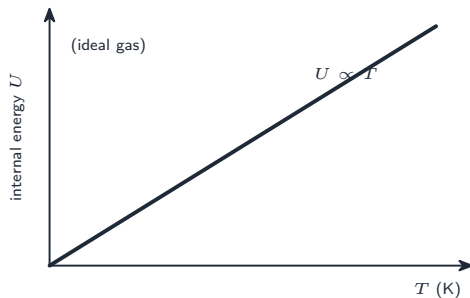
Each ideal-gas molecule carries $\frac{3}{2}kT$ on average (Topic 15), and there is no molecular PE, so

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

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

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

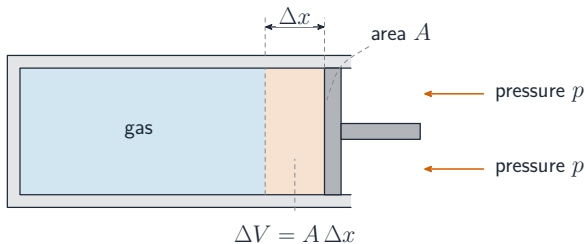
Temperature and internal energy, in detail



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



Work done by a gas

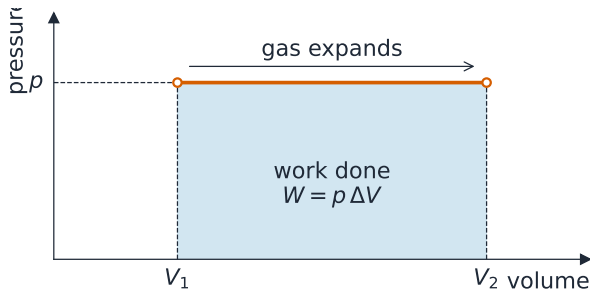


A gas pushing a **piston** of area A out by Δx sweeps $\Delta V = A\Delta x$. At constant pressure the work is

$$W = p\Delta V$$

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, $p\Delta V$.

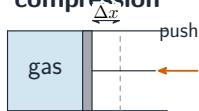
Gas at $1.0 \times 10^5 \text{ Pa}$ expands from 2.0×10^{-3} to $5.0 \times 10^{-3} \text{ m}^3$. Find the work done *by* it.

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

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

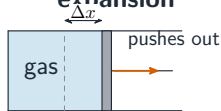
The sign convention

compression



$\Delta V < 0$, so $W > 0$

expansion



$\Delta V > 0$, so $W < 0$

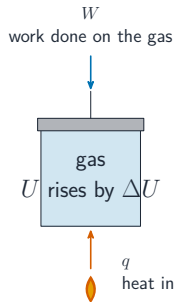
This syllabus writes $\Delta U = q + W$ with W the work done **on** the gas:

- **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 , 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

ΔU is fixed by the change of state (for an ideal gas, by the change in temperature).

- all heat, no work: $\Delta U = q$ (constant-volume heating).
- all work, no heat: $\Delta U = W$ (insulated compression or expansion).

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.

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

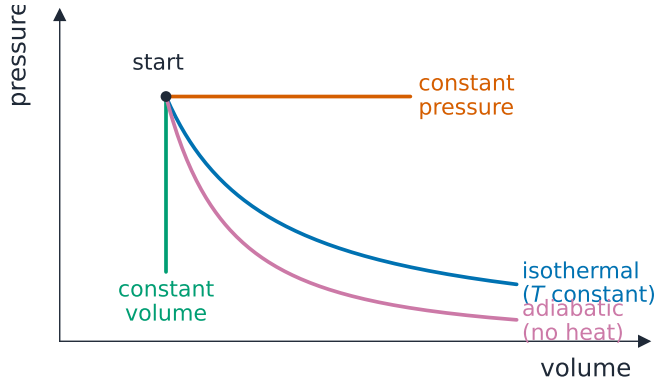
ΔU is fixed by the **change of state** (for an ideal gas, by ΔT alone). The **mix** of q and W that delivers it is up to the process.

The four standard processes

Process	fixed	ΔU	then
Isothermal	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	$q = 0$	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 :

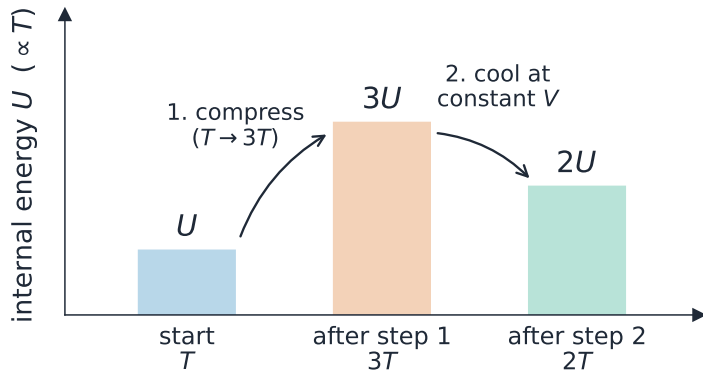
1. $U \rightarrow 3U$, so $\Delta U_1 = 2U$, $W_1 = +W$

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

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 $2T$

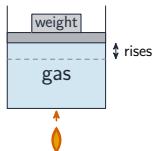
Constant volume vs constant pressure

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.

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** 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.

$Q = \Delta U + W$ still holds. A constant temperature means constant *kinetic* energy, never constant internal energy.

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