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

1 Internal energy

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.

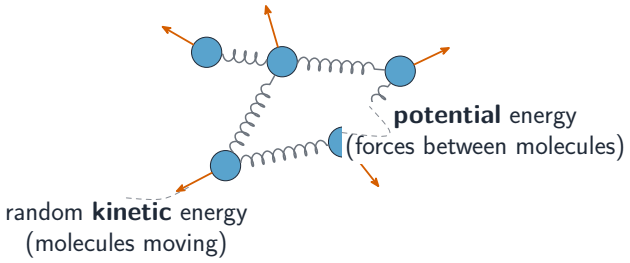
random KE

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

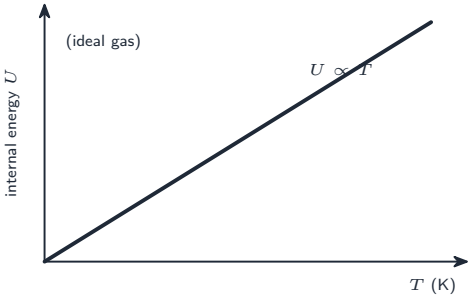
Each ideal-gas molecule carries $\frac{3}{2}kT$ on average (Topic 15), and there is no molecular PE, so

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

$U \propto T$: 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

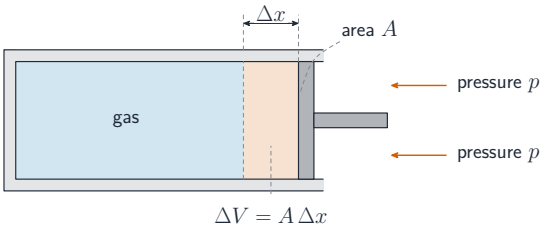


For an ideal gas $U \propto T$ – a straight line through the origin. A phase change breaks it: U climbs while T stands still.

2

Work

Work done by a gas

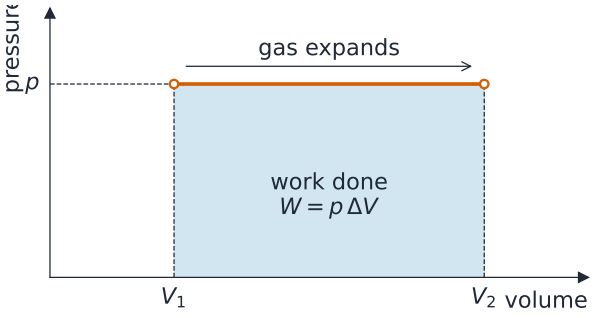


A gas pushing a **piston** 活塞 of area A out by Δx sweeps $\Delta V = A\Delta x$. At constant pressure the work is

$$W = p\,\Delta V$$

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



work = area under p–V · $W = \int p\,dV$

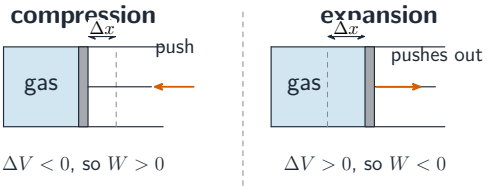
At constant pressure the area under the line is a rectangle, $p\Delta V$.

Gas at $1.0 \times 10^5\,\mathrm{Pa}$ expands from 2.0×10^{-3} to $5.0\times10^{-3}\,\mathrm{m}^3$. Find the work done *by* it.

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

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

The sign convention



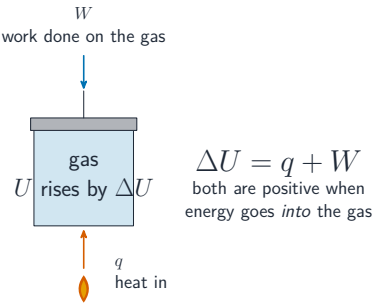
This syllabus writes $\Delta U = q + W$ with W the work done on the gas:

- 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.

3 The first law

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

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 $500\,\mathrm{J}$ of heat while doing $200\,\mathrm{J}$ of work on its surroundings. Find ΔU.

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

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

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

$\Delta U = W$ – compressing an insulated cylinder

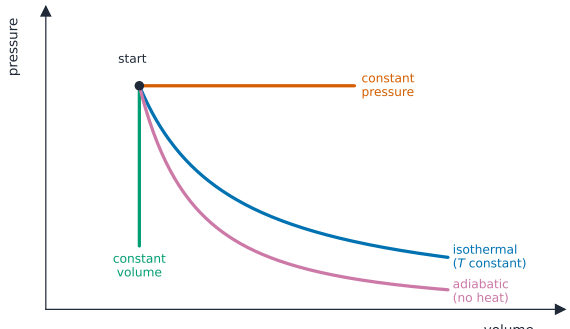
ΔU is fixed by the change of state (for an ideal gas, by ΔT alone). The mix of q and W that delivers it is up to the process.

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

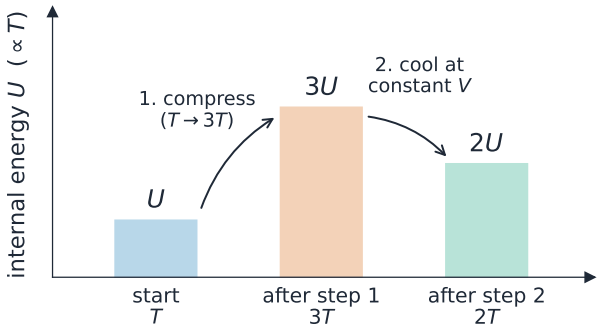
Gas at T with internal energy U: (1) compress to $3T$, work W done on it; (2) cool at constant volume to $2T$.

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

1. $U \to 3U$, 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. $T \to 2T$. Consistent.

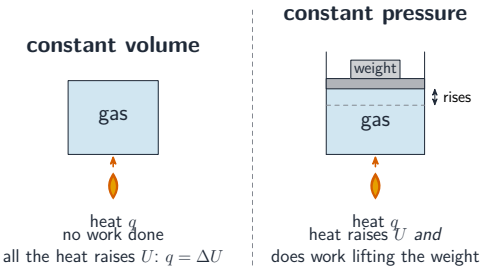
Worked example – a two-step process, in detail



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

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

Constant volume vs constant pressure



- 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** 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.

$Q = \Delta U + W$ 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).

4 Recap

Topic 16 – key takeaways

1 Internal energy

random KE + molecular PE; a *state* function

2 Ideal gas

$U = \frac{3}{2}nRT$ – depends on T 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