

Thermodynamics

A2 Physics ■ Topic 16

A-Level Physics



Thermodynamics

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

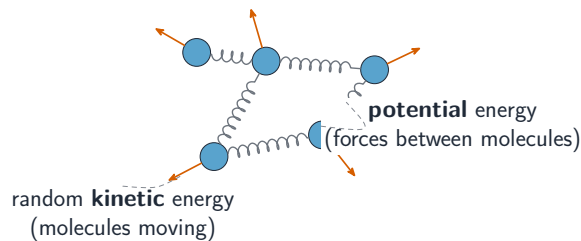
Internal energy

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



U is the total of the **random** kinetic energies and the potential energies of the molecules — not the body's motion, and not its position.

Two things U is not

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.

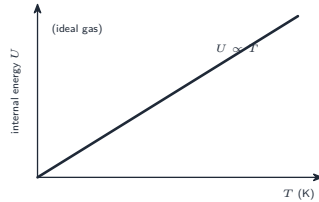
Temperature and internal energy

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

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

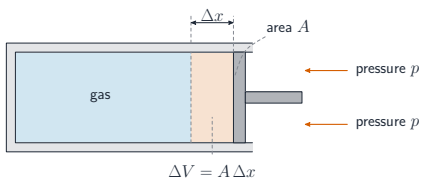
$U \propto T$: double the thermodynamic temperature, double the internal energy.

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.



2 Work

Work done by a gas

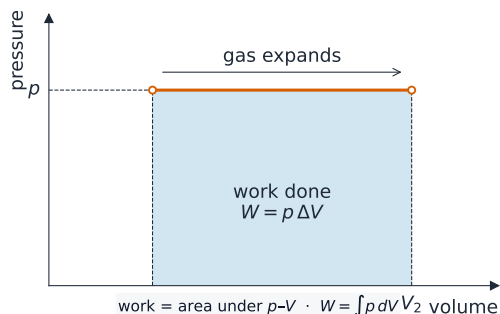


A gas pushing a **piston** 活塞 of area A out by Δx sweeps $\Delta V = A\Delta x$. At constant pressure the work is

$$W = p\Delta V$$

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



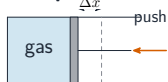
At constant pressure the area under the line is a rectangle, $p\Delta V$.

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

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

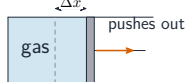
The sign convention

compression



$$\Delta V < 0, \text{ so } W > 0$$

expansion



$$\Delta V > 0, \text{ so } W < 0$$

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

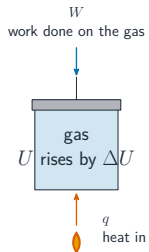
Read the question's preposition.

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

3

The first law

The first law of thermodynamics



$$\Delta U = q + W$$

both are positive when energy goes into the gas

Energy is conserved 能量守恒 when heat and work cross the 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.

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: $W = -200$ 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.

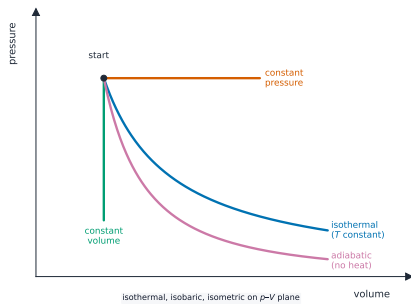
Same ΔU , different routes

All heat, no work
 $\Delta U = q$ – heating a sealed rigid container

All work, no heat
 $\Delta U = W$ – 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.

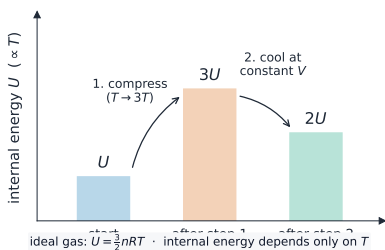
The four standard processes



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.

Worked example – a two-step process



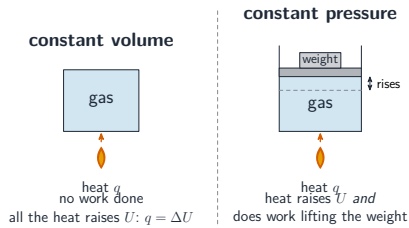
Gas at T with internal energy U : (1) compress to $3T$, work W done on it; (2) cool at constant volume to $2T$.

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

- $U \rightarrow 3U$, so $\Delta U_1 = 2U$, $W_1 = +W \Rightarrow q_1 = 2U - W$
- $\Delta U_2 = -U$, $W_2 = 0 \Rightarrow q_2 = -U$ (out)

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

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.

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Recap

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