Thermodynamics

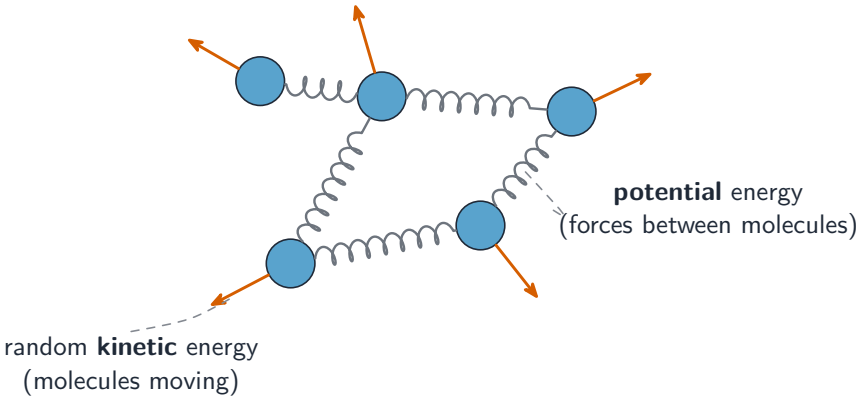
A-Level Physics

Internal energy

The **internal energy** 内能 U of a system is the sum of:

- the **random kinetic energies** 动能 of its **molecules** 分子—they fly through space (**translational** 平动 motion), and unless they are single atoms they also spin (**rotational** 转动) and shake (**vibrational** 振动), and
- the **potential energies** from the forces between the molecules.

For a real solid, liquid or gas, both parts matter. In the **ideal-gas** 理想气体 model the **intermolecular** 分子间 forces are ignored, so the molecular potential energy is zero and the internal energy is purely kinetic.



Internal energy is the molecules' random kinetic energy (the arrows) plus the potential energy of the forces between them (the springs)

Two key points:

1. U **depends only on the state** of the system (its **temperature** 温度, **pressure** 压强, **volume** 体积, **amount of substance** 物质的量) —not on the path taken to get there.
2. U **is a sum over the molecules**, not the kinetic energy of the whole object moving. A moving train of gas has bulk kinetic energy, but that is separate from U —U is the energy of the *random* molecular motion.

Temperature and internal energy

Raising an object's temperature raises the random kinetic energy of its molecules, and so raises its internal energy.

For an ideal gas every molecule has average translational kinetic energy $\frac{3}{2}kT$ (Topic 15). With zero intermolecular potential energy, the total internal energy is

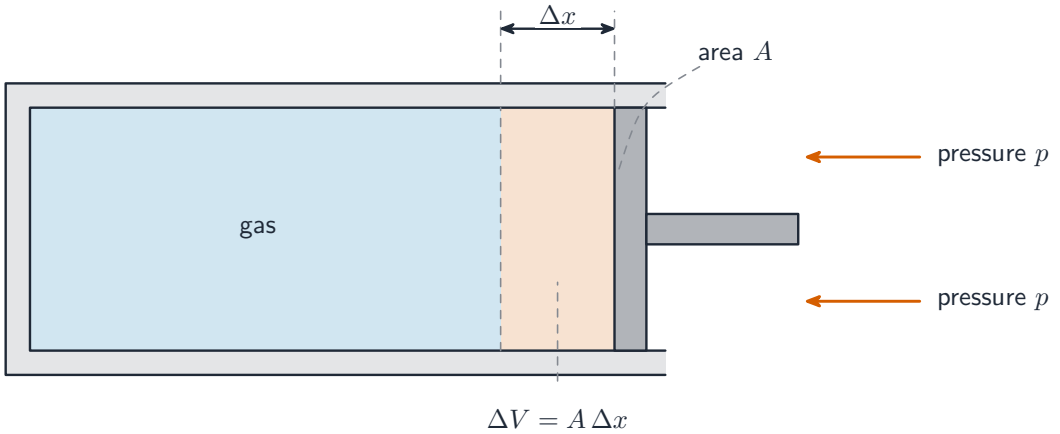
$$U = \tfrac{3}{2}NkT = \tfrac{3}{2}nRT.$$

So the internal energy of an ideal gas is directly proportional to the **thermodynamic temperature** 热力学温度. Doubling T doubles U. This is **only** exact for an ideal gas.

During a **phase change** 相变 (melting or boiling) of a real substance, U rises because the molecular potential energy rises (bonds breaking), even though the temperature stays constant.

Work done on or by a gas

A steam turbine does work as expanding steam pushes its blades around.



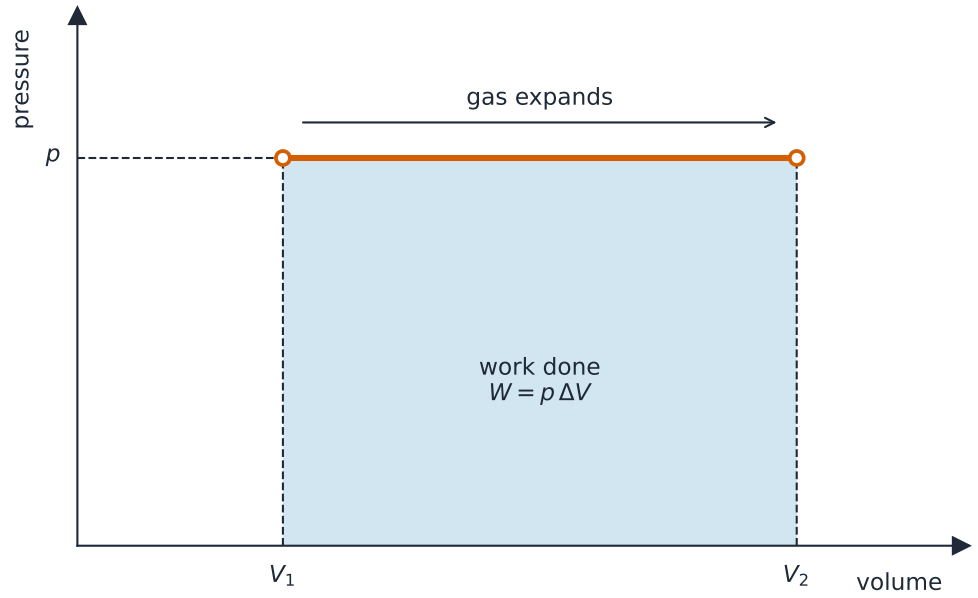
A gas pushing a piston of area A out by Δx does work $W = p\,\Delta V$ on the surroundings (swept volume $\Delta V = A\,\Delta x$)

When a gas changes volume against an outside pressure, mechanical work is done. At constant pressure p with a small volume change ΔV, the size of the work is

$$W = p\Delta V.$$

Worked example. A gas at a constant pressure of 1.0×10^5 Pa expands from 2.0×10^{-3} m^3 to 5.0×10^{-3} m^3. Find the work done by the gas.

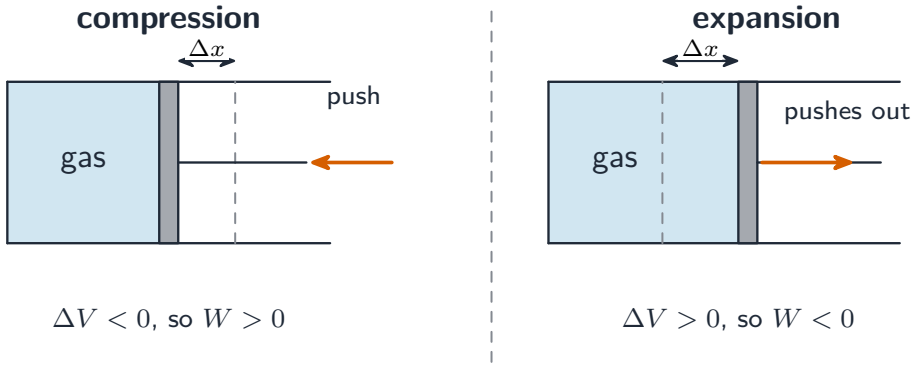
$$W = p\,\Delta V = (1.0 \times 10^5)(5.0 \times 10^{-3} - 2.0 \times 10^{-3}) = 300 \text{ J}.$$



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

Sign convention in this syllabus

This syllabus writes the first law as $\Delta U = q + W$, where W is the work done **on** the gas and q is the energy put in **by heating**.



Work done on the gas: positive when it is compressed, negative when it expands

- when the gas is **compressed**, ΔV is negative and the work done **on** the gas is positive —the gas gains energy.
- when the gas **expands**, ΔV is positive and the work done **on** the gas is negative — the gas loses energy (it does work on the surroundings).

Watch which form a question wants:

- "work done **on** the gas" —positive when compressing.
- "work done **by** the gas" —the opposite sign, positive when expanding.

At constant volume ($\Delta V = 0$), no work is done.

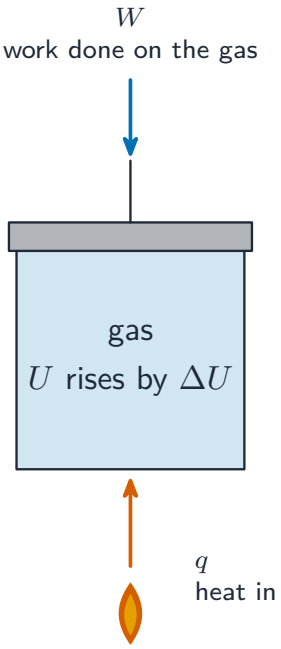
First law of thermodynamics

A power station is a heat engine: it converts heat into useful work.

The **first law of thermodynamics** 热力学第一定律 says that **energy is conserved when heat and work pass between a system and its surroundings**:

$$\Delta U = q + W,$$

where ΔU is the rise in internal energy, q is the energy added **by heating** (positive in, negative out), and W is the work done **on** the gas (positive when compressed). This is **conservation of energy** 能量守恒 for a gas.



$$\Delta U = q + W$$

both are positive when
energy goes *into* the gas

Both heating (q) and work done on the gas (W) put energy in, raising the internal energy by ΔU

Worked example. A gas absorbs 500 J of heat while it expands and does 200 J of work on its surroundings. Find the change in its internal energy.

The gas does work, so the work done *on* it is $W = -200$ J:

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

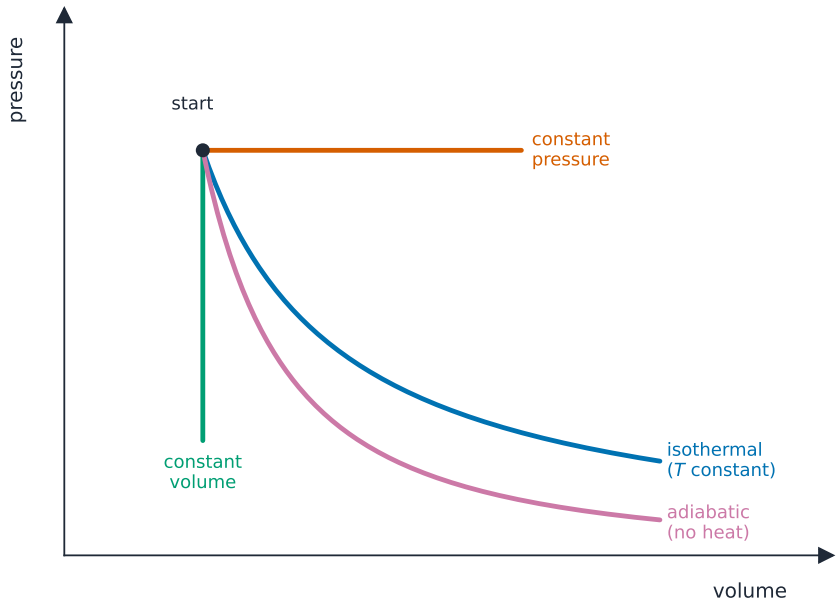
Reading the equation

ΔU is fixed by the change of state (for an ideal gas, by the change in temperature). The same ΔU can come from different mixes of q and W:

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

Standard processes

For an ideal gas, $\Delta U = \frac{3}{2} nR\Delta T$ —it depends only on ΔT.



The four standard processes, all starting from the same state

Process	**What stays constant**	ΔU	W **(on gas)**	q
Isothermal	T	0	W	$-W$
Constant volume	V	$\frac{3}{2}nR\Delta T$	0	ΔU
Constant pressure	p	$\frac{3}{2}nR\Delta T$	$-p\Delta V$	$\Delta U - W$
Adiabatic	(no heat)	varies	W	0

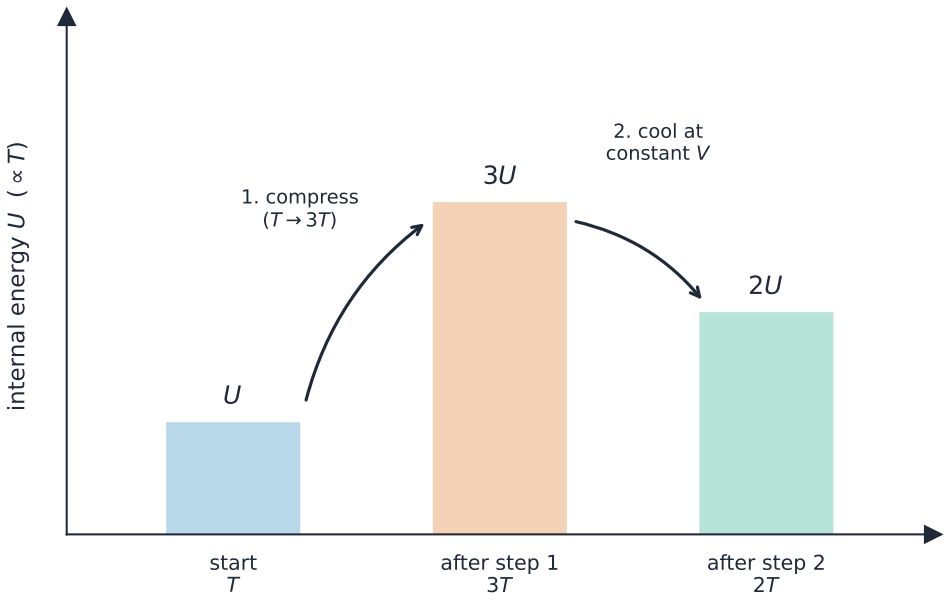
Read each row with the first law $\Delta U = q + W$:

- **Isothermal** 等温 (constant T): $\Delta T = 0$, so $\Delta U = 0$. Then $q = -W$ —any heat that goes in comes straight back out as work.
- **Adiabatic** 绝热 (no heat flow): $q = 0$, so $\Delta U = W$. The gas warms up only because work is done on it.
- **Constant volume** (sealed rigid container): no work is done ($\Delta V = 0$), so all the heat goes into internal energy: $q = \Delta U$.
- **Constant pressure** (gas pushing a **piston** 活塞): the gas does work as it expands, so the heat you supply does two jobs —it raises the internal energy *and* does the expansion work.

Worked example: two-step process

A sample of ideal gas at temperature T with internal energy U goes through:

1. **compression** to temperature $3T$; work W is done on the gas.
2. **cooling at constant volume** to temperature $2T$.



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

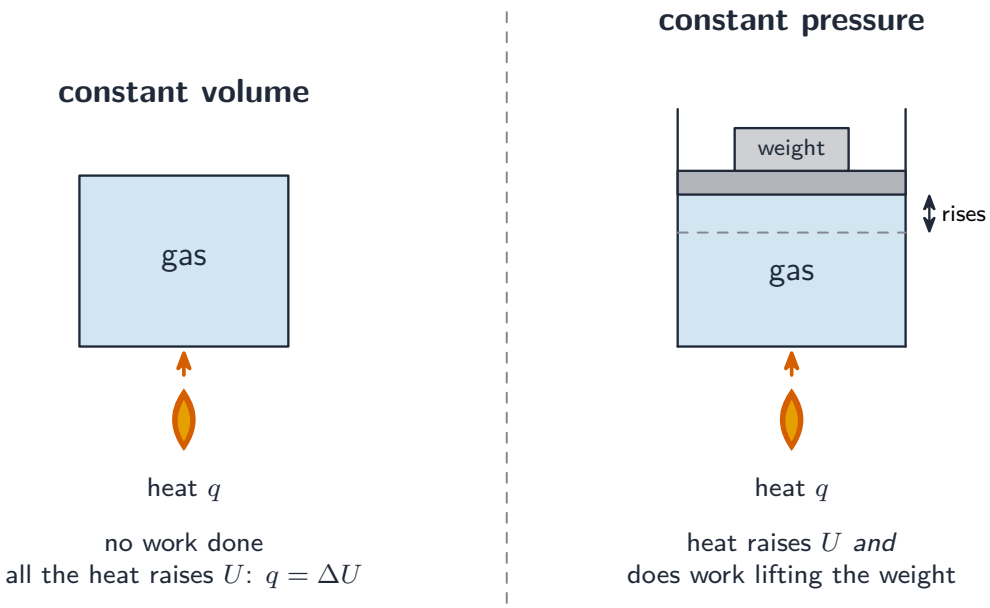
Step 1 ($T \to 3T$): $U = \frac{3}{2}nRT$, so $U \to 3U$, giving $\Delta U_1 = 2U$. $W_1 = +W$. So $q_1 = \Delta U_1 - W_1 = 2U - W$.

Step 2 ($3T \to 2T$, constant volume): $\Delta U_2 = -U$. $W_2 = 0$. So $q_2 = -U$ (heat flows out).

Check: total $\Delta U = 2U - U = U$, taking the gas from T to $2T$ ($U \to 2U$) —consistent.

Heat capacity at constant volume

For constant-volume heating of an ideal gas, $q = \Delta U = \frac{3}{2}nR\Delta T$. So the molar **heat capacity** 热容 at constant volume is $\frac{3}{2}R$ for a **monatomic** 单原子 ideal gas. (You are not required to use the symbol C_V, but the result $q = \frac{3}{2}nR\Delta T$ for constant-volume heating is.)



At constant volume all the heat raises U ($q = \Delta U$); at constant pressure the gas also does work

Heating without a temperature change

If heat is supplied during a phase change at constant pressure (e.g. boiling water), the temperature stays constant but the internal energy still rises (the **latent heat** 潜热 separates the molecules), and the gas does expansion work. The first law still holds: $\Delta U = q + W$.

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).