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 (**translational** 平动, and for non-monatomic molecules also **rotational** 转动 and **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.

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.

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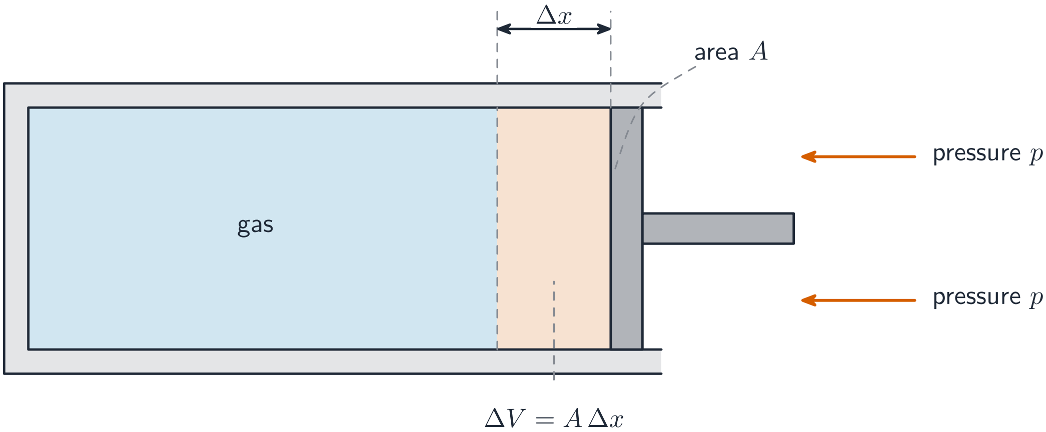

A gas pushing a piston of area A out by Δx does work $W = p\,\Delta V$ on the surroundings (swept volume $\Delta V = A\,\Delta x$)

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

$$W = p\Delta V.$$

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**.

- 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.

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.

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

In an **isothermal** 等温 process of an ideal gas, $\Delta T = 0$ so $\Delta U = 0$; then $q = -W$ (any heat in comes out as work). In an **adiabatic** 绝热 process no heat flows, so $\Delta U = W$.

For constant-volume heating (gas in a sealed rigid container), all the heat goes into internal energy: $q = \Delta U$. For a constant-pressure expansion (gas pushing a **piston** 活塞), the gas does work on the surroundings, so the heat supplied must both raise the internal energy and supply 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 $2T$.

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.)

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: $\Delta U = q + W$.