EXAM PAPERS PRACTICE

A Level Physics CIE

16. Thermodynamics

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16.1 The First Law of Thermodynamics

16.1.1 Internal energy

Defining Internal Energy

- Energy can generally be classified into two forms: kinetic or potential energy
- The molecules of all substances contain both kinetic and potential energies
- The amount of kinetic and potential energy a substance contains depends on the phases of matter (solid, liquid or gas), this is known as the **internal energy**
- The internal energy of a substance is defined as:

The sum of the random distribution of kinetic and potential energies within a system of molecules

- The symbol for internal energy is U, with units of **Joules** (J)
- The internal energy of a system is determined by:
  - Temperature
  - The random motion of molecules
  - The phase of matter: gases have the highest internal energy, solids have the lowest
- The internal energy of a system can increase by:
  - Doing work on it
  - Adding heat to it
- The internal energy of a system can decrease by:
  - Losing heat to its surroundings

Exam Tip

When an exam question asks you to define "internal energy", you can lose a mark for not mentioning the "**random motion**" of the particles or the "**random distribution**" of the energies, so make sure you include one of these in your definition!

Internal Energy & Temperature

- The internal energy of an object is intrinsically related to its temperature
- When a container containing gas molecules is heated up, the molecules begin to **move around faster,** increasing their kinetic energy
- If the object is a solid, where the molecules are tightly packed, when heated the molecules begin to **vibrate** more
- Molecules in liquids and solids have both kinetic and potential energy because they are close together and bound by intermolecular forces
- However, ideal gas molecules are assumed to have **no intermolecular forces**
    - This means there have **no potential energy, only kinetic energy**

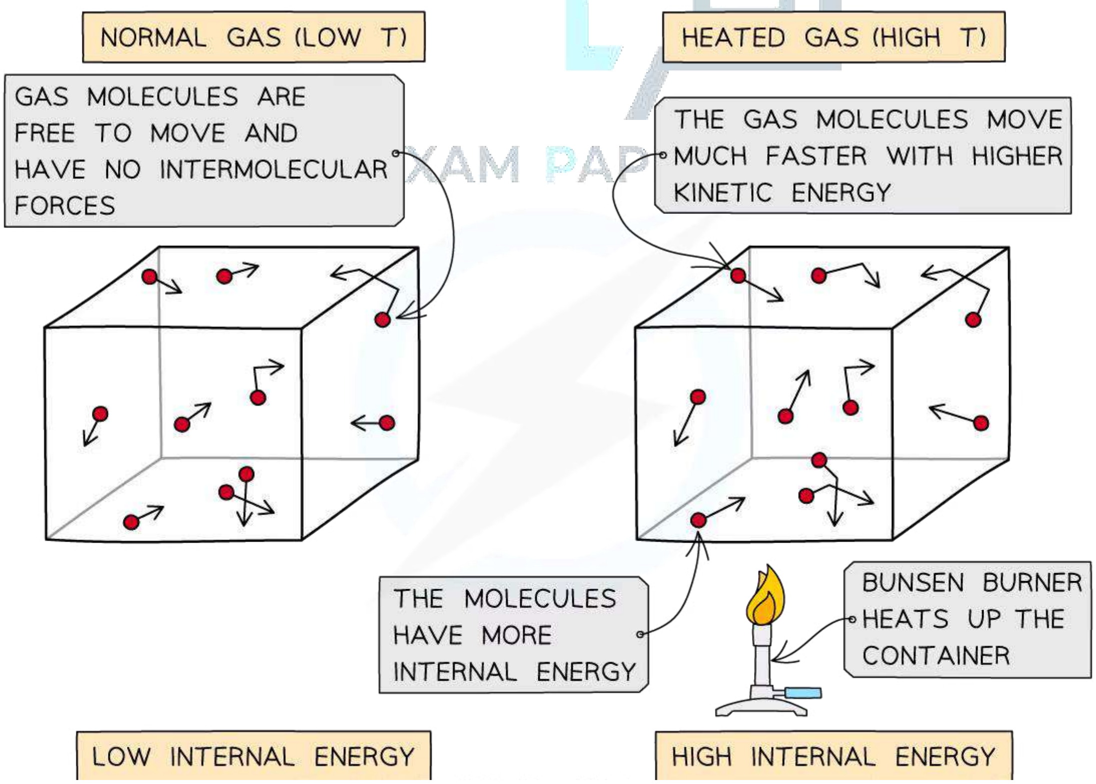
- The (change in) internal energy of an ideal gas is equal to:

$$\Delta U = \frac{3}{2} k\Delta T$$

- Therefore, the change in internal energy is proportional to the change in temperature:

$$\Delta U \propto \Delta T$$

- Where:
    - ΔU = change in internal energy (J)
    - ΔT = change in temperature (K)



As the container is heated up, the gas molecules move faster with higher kinetic energy and therefore higher internal energy

Worked Example

A student suggests that, when an ideal gas is heated from 50 ^oC to 150 ^oC, the internal energy of the gas is trebled.State and explain whether the student's suggestion is correct.

Step 1:

Write down the relationship between internal energy and temperature

The internal energy of an ideal gas is directly proportional to its temperature

$$\Delta U \propto \Delta T$$

Step 2:

Determine whether the change in temperature (in K) increases by three times

The temperature change is the **thermodynamic** temperature ie. Kelvin

The temperature change in degrees from 50 ^oC to 150 ^oC increases by three times

The temperature change in Kelvin is:

$$50\ ^oC + 273.15 = 323.15\ K$$

$$150\ ^oC + 273.15 = 423.15\ K$$

$$\frac{423.15}{323.15} = 1.3$$

Therefore, the temperature change, in Kelvin, does **not** increase by three times

Step 3:

Write a concluding statement relating the temperature change to the internal energy

The internal energy is directly proportional to the temperature

The thermodynamic temperature has not trebled, therefore, neither has the internal energy

Therefore, the student is incorrect

Exam Tip

If an exam question about an ideal gas asks for the **total internal energy**, remember that this is equal to the **total kinetic energy** since an ideal gas has **zero potential energy**

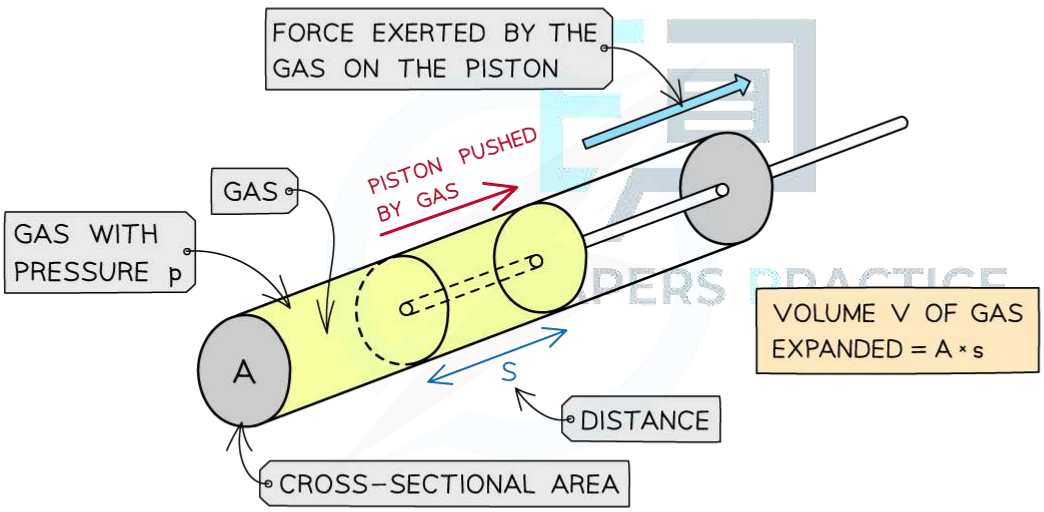
Work Done by a Gas

- When a gas expands, it does work on its surroundings by exerting pressure on the walls of the container it's in
- This is important, for example, in a steam engine where expanding steam pushes a piston to turn the engine
- The work done when a volume of gas changes at constant pressure is defined as:

$$W = p\Delta V$$

- Where:
  - W = work done (J)
  - p = external pressure (Pa)
  - V = volume of gas (m^3)

- For a gas inside a cylinder enclosed by a moveable piston, the force exerted by the gas pushes the piston outwards
- Therefore, the gas **does work on the piston**



The gas expansion pushes the piston a distance s

Derivation

- The volume of gas is at constant pressure. This means the force *F* exerted by the gas on the piston is equal to :

$$F = p \times A$$

- Where:
  - *p* = pressure of the gas (Pa)
  - *A* = cross-sectional area of the cylinder (m^2)

- The definition of work done is:

$$W = F \times s$$

- Where:
  - F = force (N)
  - s = displacement in the direction of force (m)
- The displacement of the gas d multiplied by the cross-sectional area A is the increase in volume ΔV of the gas:

$$W = p \times A \times s$$

- This gives the equation for the work done when the volume of a gas changes at constant pressure:

$$W = p\Delta V$$

- Where:
  - ΔV = increase in the volume of the gas in the piston when expanding (m^3)

- This is assuming that the surrounding pressure p does not change as the gas expands
- This will be true if the gas is expanding against the pressure of the atmosphere, which changes very slowly
- When the gas **expands** (V increases), work is done **by** the gas
- When the gas is **compressed** (V decreases), work is done **on** the gas

Worked Example

When a balloon is inflated, its rubber walls push against the air around it.Calculate the work done when the balloon is blown up from 0.015 m^3 to 0.030 m^3.Atmospheric pressure = 1.0×10^5 Pa.

Step 1: Write down the equation for the work done by a gas

$$W = p\Delta V$$

Step 2: Substitute in values

$$\Delta V = \text{final volume} - \text{initial volume} = 0.030 - 0.015 = 0.015 \text{ m}^3$$

$$W = (1.0 \times 10^5) \times 0.015 = 1500 \text{ J}$$

Exam Tip

The pressure p in the work done by a gas equation is not the pressure of the gas but the pressure of the surroundings. This is because when a gas expands, it does work **on** the surroundings.

The First Law of Thermodynamics

- The first law of thermodynamics is based on the principle of conservation of energy
- When energy is put into a gas by heating it or doing work on it, its internal energy must increase:

    The increase in internal energy = Energy supplied by heating + Work done on the system

- The first law of thermodynamics is therefore defined as:

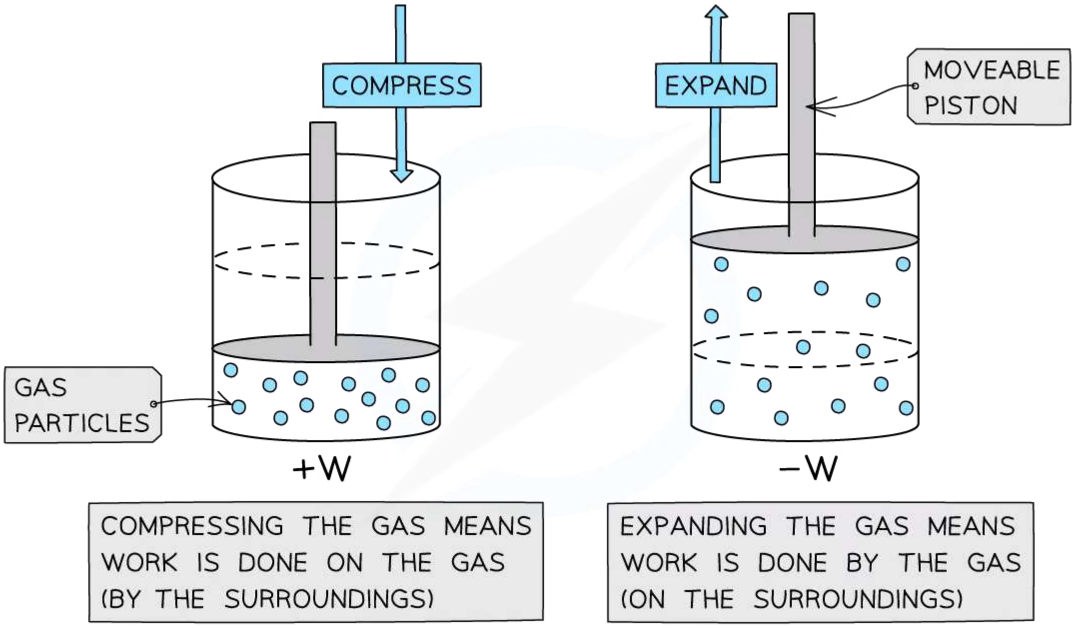
$$\Delta U = q + W$$

- Where:
    - ΔU = increase in internal energy (J)
    - q = energy supplied to the system by heating (J)
    - W = work done on the system (J)
- The first law of thermodynamics applies to **all** situations, not just for gases
    - There is an important sign convention used for this equation
- A **positive** value for internal energy ($+\Delta U$) means:
    - The internal energy ΔU **increases**
    - Heat q is **added** to the system
    - Work W is done **on** the system (or **by** a gas)
- A **negative** value for internal energy ($-\Delta U$) means:
    - The internal energy ΔU **decreases**
    - Heat q is **taken away** from the system
    - Work W is done **by** the system (or **on** a gas)
- This is important when thinking about the expansion or compression of a gas
- When the gas **expands**, it transfers some energy (does work) to its surroundings
- This decreases the overall energy of the gas
- Therefore, when the gas expands, work is done **by** the gas ($-W$)

    When a gas expands, work done W is negative

- When the gas is **compressed**, work is done **on** the gas ($+W$)

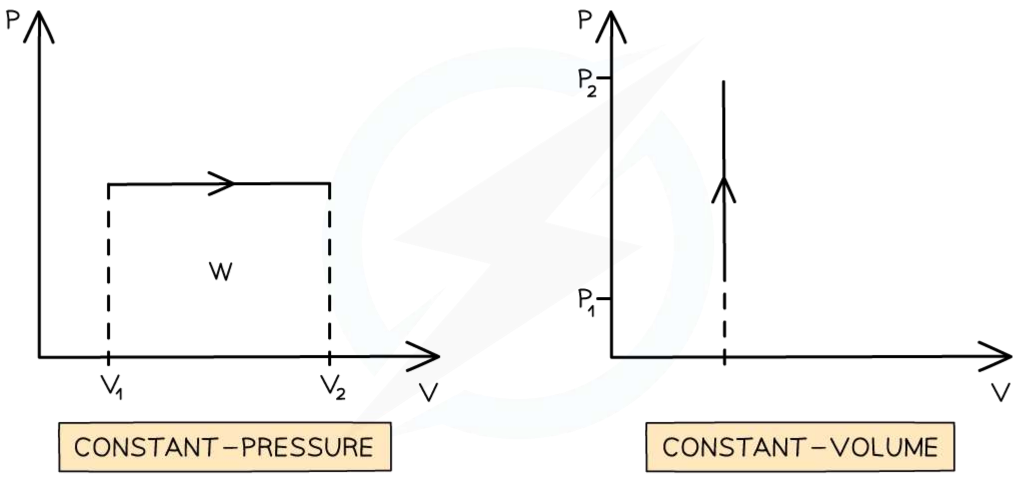
    When a gas is compressed, work done W is positive



Positive or negative work done depends on whether the gas is compressed or expanded

Graphs of Constant Pressure & Volume

- Graphs of pressure p against volume V can provide information about the work done and internal energy of the gas
  - The work done is represented by the area under the line
- A constant pressure process is represented as a **horizontal line**
  - If the volume is increasing (expansion), work is done **by** the gas and internal energy increases
  - If the arrow is reversed and the volume is decreasing (compression), work is done **on** the gas and internal energy decreases
- A constant volume process is represented as a **vertical line**
  - In a process with constant volume, the area under the curve is **zero**
  - Therefore, no work is done when the volume stays the same



Work is only done when the volume of a gas changes

Worked Example

The volume occupied by 1.00 mol of a liquid at 50 °C is 2.4×10^{-5} m^3. When the liquid is vaporised at an atmospheric pressure of 1.03×10^5 Pa, the vapour has a volume of 5.9×10^{-2} m^3. The latent heat to vaporise 1.00 mol of this liquid at 50 °C at atmospheric pressure is 3.48×10^4 J. Determine for this change of state the increase in internal energy ΔU of the system.

Step 1: Write down the first law of thermodynamics

$$\Delta U = q + W$$

Step 2: Write the value of heating *q* of the system

This is the latent heat, the heat required to vaporise the liquid = **3.48×10^4 J**

Step 3: Calculate the work done *W*

$$W = p\Delta V$$

ΔV = final volume − initial volume = **$5.9 \times 10^{-2} - 2.4 \times 10^{-5} = 0.058976$ m^3**

p = atmospheric pressure = **1.03×10^5 Pa**

W = **$(1.03 \times 10^5) \times 0.058976 = 6074.528 = 6.07 \times 10^3$ J**

Since the gas is expanding, this work done is negative

$$W = -6.07 \times 10^3 \text{ J}$$

Step 4: Substitute the values into first law of thermodynamics

$$\Delta U = 3.48 \times 10^4 + (-6.07 \times 10^3) = 28\,730 = 29\,000 \text{ J (2 s.f.)}$$