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6.4 Thermal Energy Transfer

PHYSICS

AQA A Level Revision Notes

A Level Physics AQA

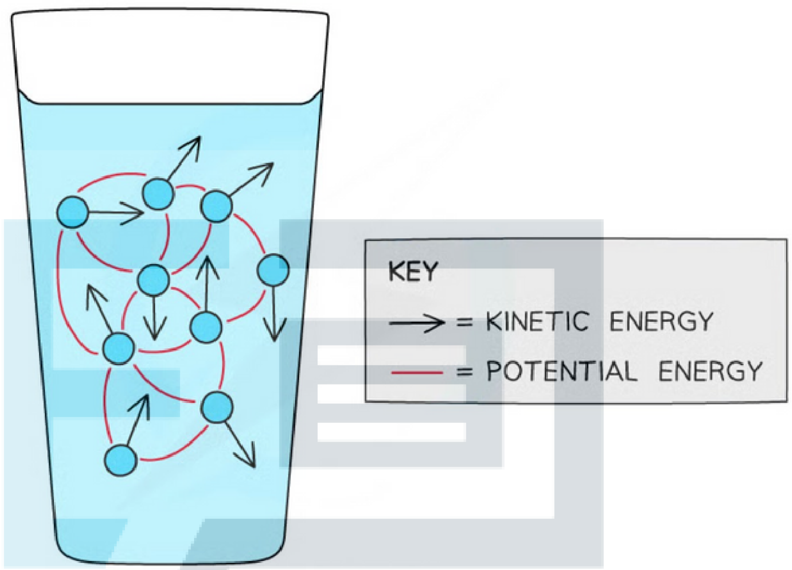
6.4 Thermal Energy Transfer

CONTENTS

6.4.1 Internal Energy

Internal Energy

- Energy can be classified into two forms: kinetic or potential energy
- The molecules of all substances contain both kinetic and potential energies
  - Kinetic energy is due to the speed of the molecules
  - Potential energy is due to the separation between the molecules



- The amount of kinetic and potential energy a substance contains depends on its phase 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 randomly distributed kinetic and potential energies of the particles in a body

- The symbol for internal energy is *U*, with units of **Joules** (J)
- Particles are randomly distributed, meaning they all have different speeds and separations
- The internal energy of a system is determined by:
  - **Temperature** (higher temperature, higher kinetic energy and vice versa)
  - The **random motion** of molecules
  - The **phase of matter**: gases have the highest internal energy, solids have the lowest
  - **Intermolecular forces** between the particles (greater intermolecular forces, higher potential energy and vice versa) - this is linked to the phase (solid, liquid, gas) that the matter is in
- 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
  - **Changing state** from a solid to a liquid or liquid to a gas

Exam Tip

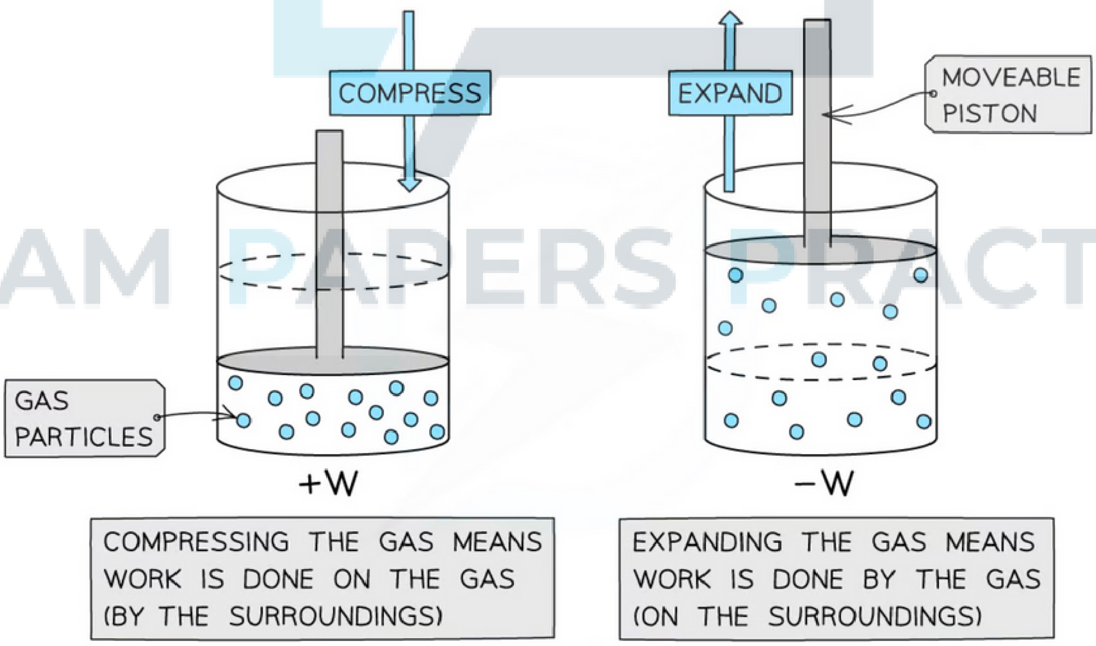
When an exam question asks you to define "internal energy", you can lose a mark for not mentioning the "**random motion**" of the particles or the "**random distribution**" of the energies, so make sure you include one of these in your definition!

The First Law of Thermodynamics

- The First Law of Thermodynamics, which is based on the principle of conservation of energy, states:

  The internal energy of a system is increased when energy is transferred to it by heating or when work is done on it (and vice versa)

- This is important when thinking about the **expansion** or **compression** of a gas
  - The First Law of Thermodynamics applies to **all** situations, not just for gases
- When a gas **expands** (its volume increases), work is done **by** the gas **on** the surroundings
  - This decreases the internal energy of the gas
- When a gas is **compressed** (its volume decreases), work is done **on** the gas **by** the surroundings
  - This increases the internal energy of the gas
- The 'gas' is sometimes referred to as the 'system'
- A gas can have work done **on** or **by** it when it is in a cylindrical container with a moveable piston



Work is done on the gas when it is compressed and by the gas when it expands

- When the piston moves down the cylinder, it compresses the gas molecules (work is done **on** the gas)
  - The molecules are pushed closer together
  - Therefore, they have **higl**

  - This **increases** the internal energy of the gas

- When the piston moves up the cylinder, it expands the gas molecules (work is done **by** the gas)
  - The molecules are spread further apart
  - Therefore, they have **lower** kinetic energy as they move slower
  - This **decreases** the internal energy of the gas

- The same increase in internal energy can be achieved by not doing work (i.e. no expansion or contraction), but by **heating** the gas instead
  - Increasing the temperature of the gas means the molecules move around **faster**
  - They, therefore, have **higher** kinetic energy and increased internal energy

Exam Tip

Remember that the number of molecules in the container **always** remains the **same** whether the gas is expanding or contractingTry not to get too hung up on 'positive' and 'negative' work, this is more relevant for the 'Engineering Physics' option module. The important ideas to remember are when work is done **by** or **on** the gas (or system) and the ways in which internal energy changes.

6.4.3 Specific Heat Capacity

Specific Heat Capacity

- When a substance is heated, its temperature rises causing the particles within it to gain kinetic energy
  - The amount of energy required to raise the temperature of a substance is given by its **specific heat capacity**
- The specific heat capacity of a substance is defined as:

  The amount of thermal energy required to raise the temperature of 1 kg of a substance by 1 °C (or 1 K) without a change of state

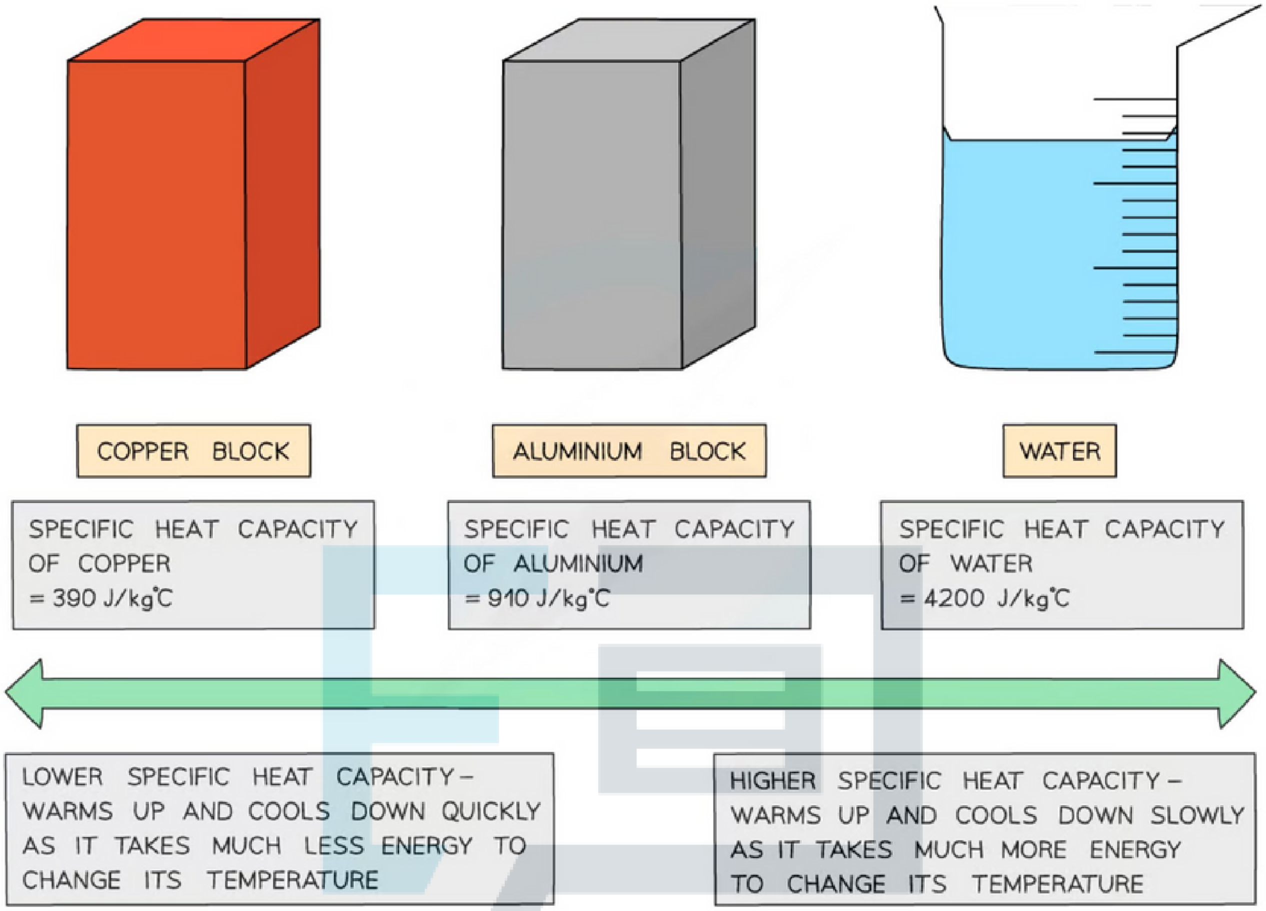
- This quantity determines the amount of energy needed to change the temperature of a substance
- Specific heat capacity has the symbol **c** and is measured in units of **Joules per kilogram per Kelvin** ($J\ kg^{-1} K^{-1}$) or **Joules per kilogram per Celsius** ($J\ kg^{-1}\ °C^{-1}$)
  - Different substances have different specific heat capacities
  - Specific heat capacity is mainly used for liquids and solids
- From the definition of specific heat capacity, it follows that:
  - The heavier the material, the more thermal energy required to raise its temperature
  - The larger the change in temperature, the higher the thermal energy required to achieve this change

Calculating Specific Heat Capacity

- The amount of thermal energy Q needed to raise the temperature by $\Delta\theta$ for a mass *m* with specific heat capacity c is equal to:

$$\Delta Q = mc\Delta\theta$$

- Where:
  - ΔQ = change in thermal energy (J)
  - *m* = mass of the substance you are heating up (kg)
  - c = specific heat capacity of the substance ($J\ kg^{-1} K^{-1}$ or $J\ kg^{-1}\ °C^{-1}$)
  - $\Delta\theta$ = change in temperature (K or °C)



Low v high specific heat capacity

- If a substance has a **low** specific heat capacity, it heats up and cools down quickly
- If a substance has a **high** specific heat capacity, it heats up and cools down slowly
- The specific heat capacity of different substances determines how useful they would be for a specific purpose eg. choosing the best material for kitchen appliances

Table of values of specific heat capacity for various substances

Substance	Specific heat capacity / $J\,kg^{-1}K^{-1}$
Aluminium	910
Copper	390
Lead	126
Glass	500 – 680
Water	4200
Mercury	140

- Good electrical conductors, such as copper and lead, are **excellent conductors** of hea due to their **low** specific heat capacity

Worked Example

A kettle is rated at 1.7 kW. A mass of 650 g of a liquid at 25 °C is poured into a kettle. When the kettle is switched on, it takes 3.5 minutes to start boiling. Calculate the specific heat capacity of the liquid.

Step 1: Calculate the Energy from the power and time

$$\text{Energy} = \text{Power} \times \text{Time}$$

$$\text{Power} = 1.7\text{ kW} = 1.7 \times 10^3\text{ W}$$

$$\text{Time} = 3.5\text{ minutes} = 3.5 \times 60 = 210\text{ s}$$

$$\text{Energy} = 1.7 \times 10^3 \times 210 = 3.57 \times 10^5\text{ J}$$

Step 2: Thermal energy equation

$$\Delta Q = mc\Delta\theta$$

Step 3: Rearrange for specific heat capacity

$$c = \frac{\Delta Q}{m\Delta\theta}$$

Step 4: Substitute in values

$$\Delta\theta = 100 - 25 = 75°C$$

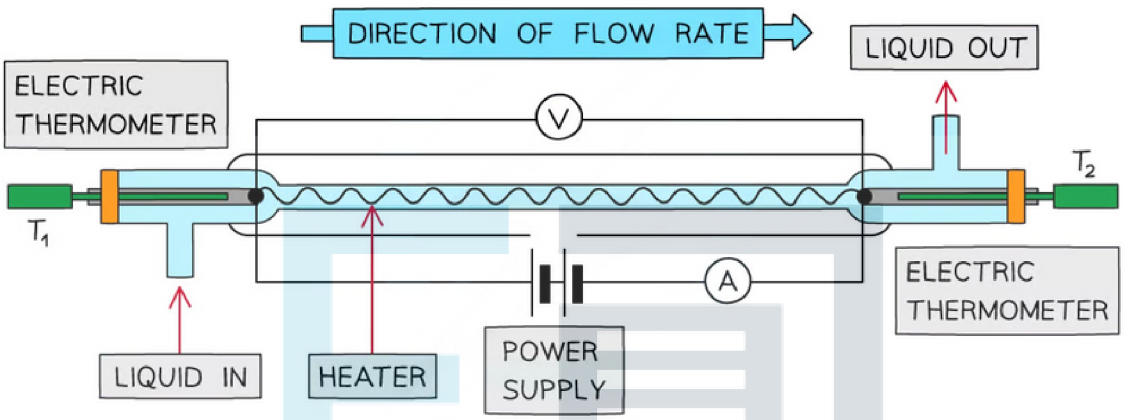
$$c = \frac{3.57 \times 10^5}{650 \times 10^{-3} \times 75} = 7323.07... = 7300 \text{ J kg}^{-1}\,°C^{-1} \text{ (2 s.f)}$$

Exam Tip

The difference in temperature Δθ will be exactly the same whether the temperature is given in Celsius or Kelvin. Therefore, there is no need to convert between the two since the **difference** in temperature will be the same for both units.

Continuous Flow

- The specific heat capacity of a fluid can be found using a continuous-flow calorimeter
- A fluid flows continuously over a heating element where energy is transferred to the fluid
  - It is assumed that the heat transferred from the apparatus to the surroundings is constant
- For this experiment, the flow rate and the potential difference is changed, keeping the change in temperature of the fluid constant



A continuous-flow calorimeter

- A fluid flows through an electrical heating wire. The rise in temperature of the fluid is measured using the electric thermometers and is calculated by:

$$\Delta\theta = T_2 - T_1$$

- To find the mass of the fluid, the flow rate is recorded and multiplied by the time taken t to give the mass of the fluid that flows in as m_1
- The current I and potential difference V are also recorded
- The flow rate is then altered to give a mass m_2 and the potential difference of the power supply is changed so the temperature difference, $\Delta\theta$ stays the same
- The specific heat capacity is found by assuming the thermal losses to the surroundings are constant for both flow rates

- For the first flow rate, the electrical energy supplied to the fluid in time t_1 is:

$$I_1V_1t_1 = Q_1 = m_1c\Delta\theta + E_{lost}$$

- Where E_{lost} is the thermal energy lost to the surroundings

- The second flow rate is:

$$I_2V_2t_2 = Q_2 = m_2c\Delta\theta + E_{lost}$$

- Since E_{lost} is assumed to be the same, subtracting the first flow rate equation from the second gives the equation:

$$I_2V_2t_2 - I_1V_1t_1 = Q_2 - Q_1 = (m_2 - m_1)c\Delta\theta$$

- Rearranging this for the specific heat capacity of the fluid, c gives the final equation:

$$c = \frac{Q_2 - Q_1}{(m_2 - m_1)\Delta\theta}$$

Worked Example

Calculate the specific heat capacity of a liquid using the following data measured in two experiments using the continuous flow method:

- Time of each experiment = 40 s
- T_1 in both experiments = 15 °C
- T_2 in both experiments = 4 °C
- p.d across the heater in experiment 1, V_1 = 14.0 V
- p.d across the heater in experiment 2, V_2 = 9.0 V
- Current through heater in both experiments = 3.0 A
- Mass of water flowing in experiment 1, m_1 = 136.0 g
- Mass of water flowing in experiment 2, m_2 = 73.0 g

Step 1: Calculate the change in temperature, Δθ

$$\Delta\theta = 15 - 4 = 11°C$$

Step 2: Calculate Q_2

$$Q_2 = I_2V_2t_2 = 3.0 \times 9.0 \times 40 = 1080 \text{ J}$$

Step 3: Calculate Q_1

$$Q_1 = I_1V_1t_1 = 3.0 \times 14.0 \times 40 = 1680 \text{ J}$$

Step 4: Substitute values into the specific heat capacity equation

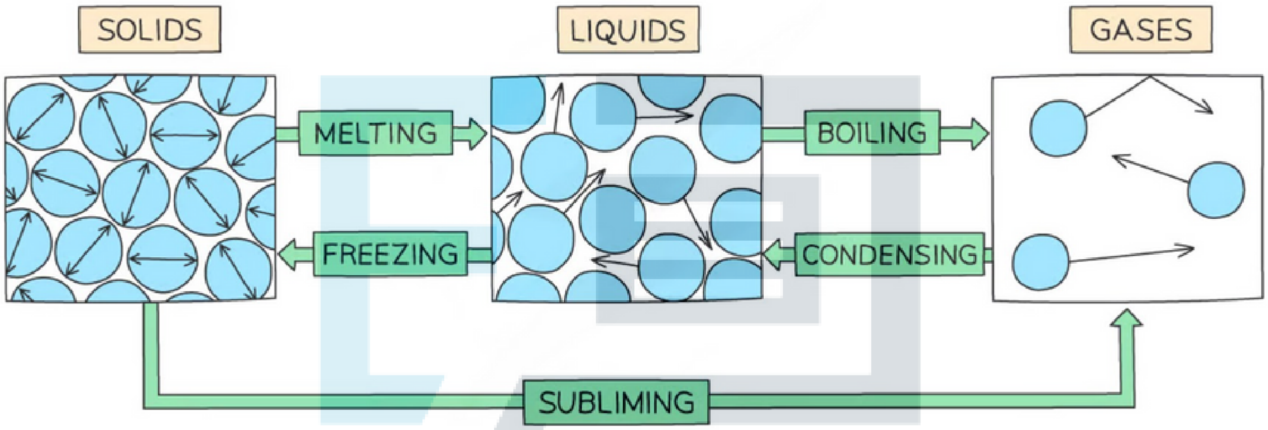
$$c = \frac{Q_2 - Q_1}{(m_2 - m_1)\Delta\theta}$$

$$c = \frac{1080 - 1680}{\left((73 \times 10^{-3}) - (136 \times 10^{-3})\right) \times 11} = \mathbf{866 \text{ J kg}^{-1}\text{ °C}^{-1}}$$

6.4.4 Latent Heat Capacity

Latent Heat Capacity

- Energy is required to change the **state** of a substance
- Examples of changes of state are:
  - **Melting** = solid to liquid
  - **Evaporation / vaporisation / boiling** = liquid to gas
  - **Sublimation** = solid to gas
  - **Freezing** = liquid to solid
  - **Condensation** = gas to liquid



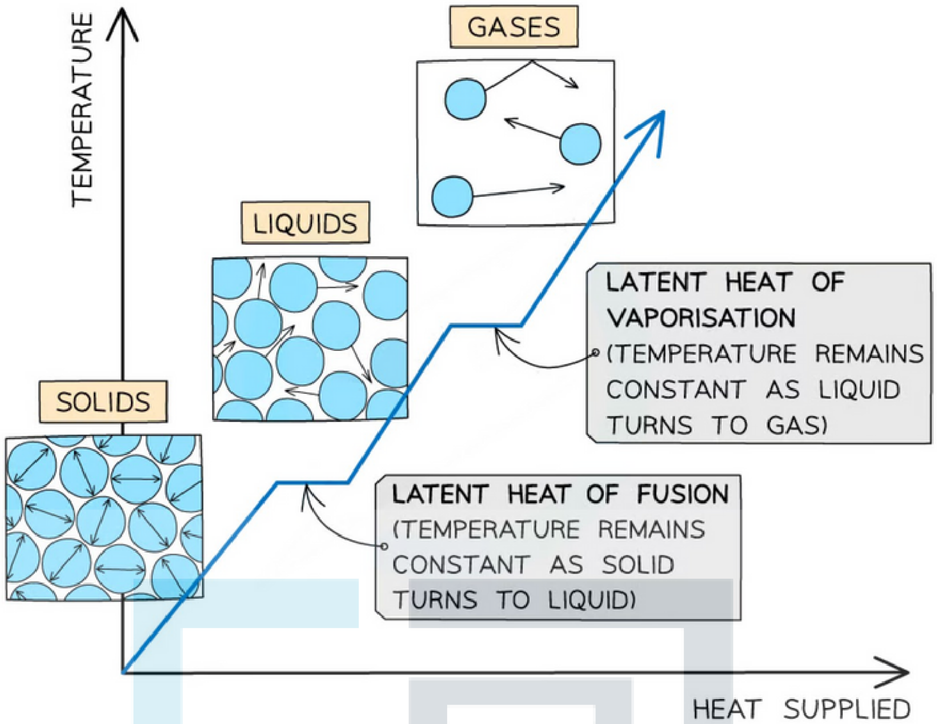
The example of changes of state between solids, liquids and gases

- When a substance changes state, there is **no temperature change**
- The energy supplied to change the state is called the **latent heat** and is defined as:

The thermal energy required to change the state of 1 kg of mass of a substance without any change of temperature

- There are two types of latent heat:
  - Specific latent heat of **fusion** (melting)
  - Specific latent heat of **vaporisation** (boiling)

- The larger the mass of the substance, the more energy will be required to change its state. Hence why specific latent heat is defined by 1 kg



The changes of state with heat supplied against temperature. There is no change in temperature during changes of state

- The horizontal line of the latent heat of fusion represents melting (if heat is supplied) or freezing (if heat is removed)
- The horizontal line of the latent heat of vaporisation represents evaporation (if heat is supplied) or condensation (if heat is removed)

- The specific latent heat of **fusion** is defined as:

  The thermal energy required to convert 1 kg of solid to liquid with no change in temperature

- Latent heat of fusion applies to:
  - Melting a solid
  - Freezing a liquid

- The specific latent heat of **vaporisation** is defined as:

  The thermal energy required to convert 1 kg of liquid to gas with no change in temperature

- Latent heat of vaporisation applies to:
  - Vaporising a liquid
  - Condensing a gas

Calculating Specific Laten

- The amount of energy Q requ

$$Q = mL$$

- Where:
  - Q = amount of thermal energy to change the state (J)
  - m = mass of the substance changing state (kg)
  - L = latent heat of fusion or vaporisation (J kg^{-1})
- The values of latent heat for water are:
  - Specific latent heat of fusion = 330 kJ kg^{-1}
  - Specific latent heat of vaporisation = 2.26 MJ kg^{-1}
- Therefore, evaporating 1 kg of water requires roughly **seven times** more energy than melting the same amount of ice to form water
- The reason for this is to do with intermolecular forces:
  - **When ice melts**: energy is required to just increase the molecule separation until they can flow freely over each other
  - **When water boils**: energy is required to completely separate the molecules until there are no longer forces of attraction between the molecules, hence this requires much more energy. Vaporisation is also doing work against atmospheric pressure
- More energy has to be supplied to separate molecules than break a solid bond, which is why the latent heat of vaporisation of water is **much greater** than the specific latent heat of fusion of water

Worked Example

The energy needed to boil a mass of 530 g of a liquid is 0.6 MJ.Calculate the specific latent heat of the liquid and state whether it is the latent heat of vaporisation or fusion.

Step 1: Write the thermal energy required to change state equation

$$Q = mL$$

Step 2: Rearrange for latent heat

$$L = \frac{Q}{m}$$

Step 3: Substitute in values

$$m = 530 \text{ g} = 530 \times 10^{-3} \text{ kg}$$

$$Q = 0.6 \text{ MJ} = 0.6 \times 10^{6} \text{ J}$$

$$L = \frac{0.6 \times 10^6}{530 \times 10^{-3}} = 1.132 \times 10^6 \text{ J kg}^{-1} = \mathbf{1.1 \text{ MJ kg}^{-1}} \text{ (2 s.f.)}$$

- L is the latent heat of vaporisation because the change in state is from liquid to gas (boiling)

Exam Tip

Use these reminders to help you remember which type of latent heat is being referred to:

- Latent heat of fusion = imagine 'fusing' the liquid molecules together to become a solid
- Latent heat of vaporisation = "water vapour" is steam, so imagine vaporising the liquid molecules into a gas

Remember to always include '**without a change in temperature**', or words to that effect, within your definitions for latent heat to gain full marks.

Energy Transfers During Phase Changes

- When a substance is heated, the molecules are given more energy in the form of kinetic and potential energy
- During a change of state (or a phase change), the key points to remember are:
  - There is no change in temperature
  - The **potential** energies of the molecules change, but **not** their **kinetic** energies
- The potential energy of the molecules is due to their separation and intermolecular bonds
  - Since they move further apart (evaporation) or closer together (condensation), their potential energy will change as a result of this
- The heat **absorbed** in melting and boiling causes the molecules to move further apart by overcoming the intermolecular forces of attraction
- The heat **released** in freezing and condensation allows the molecules to move closer together and the intermolecular forces of attraction become stronger
  - This is because the kinetic energy is proportional to the temperature
  - If there is no change in temperature, there must be no change in kinetic energy either