Section 1.1.6 – How heat affects temperature

VCAA Study Design Dot Points

- describe temperature with reference to the average translational kinetic energy of the atoms and molecules within a system:
  - explain why cooling results from evaporation using a simple kinetic energy model
- Investigate and analyse theoretically and practically the temperature required to:
  - raise the temperature of a substance: $Q = mc\Delta T$
  - change the state of a substance: $Q = mL$

When two objects at **different temperatures** are placed in contact, the temperature of the colder object **increases** and the temperature of the hotter object **decreases**.

Heat energy is transferred from the hotter object to the colder object. We call this process heating (or cooling).

The **quantity** of energy transferred depends upon:

- the **temperature** of the objects
- the **mass** of the objects
- the **material** that each object is made of

Specific Heat Capacity

Different materials have different **capacities to hold heat**.
For example, consider a 1kg sample of water and a 1 kg sample of silver. If both samples were heated to 50℃ and allowed to cool, which would reach room temperature first?



The silver sample of identical mass will cool much quicker than the water sample, as the water sample has a **greater capacity** to hold heat.

The term "**Specific heat capacity**" is defined as the **amount of energy** required to **increase the temperature of 1 kg** of the substance **by 1 °C** (or K).

NB: Water has a much higher specific heat capacity than silver.
Accordingly, it takes more energy to raise the temperature of a sample of water than silver.
Water will also retain its heat better than silver and therefore takes longer to cool.

Water has a specific heat capacity of $\mathbf{4200\ Jkg^{-1}K^{-1}}$.

By definition this means that it will take $4200\ Joules$ of energy to raise the temperature of $1\ kilogram$ of water by $1\ Kelvin$ (or 1°C).

Table.1 below displays a selection of specific heat capacities.

Substance	Specific heat capacity ($J\ kg^{-1}\ K^{-1}$)
Helium	5193
Water	4200
Human body (average)	3500
Cooking oil	2800
Ethylene glycol (used in car 'antifreeze')	2400
Ice	2100
Steam	2000
Fertile topsoil	1800
Neon	1030
Air	1003
Aluminium	897
Carbon dioxide	839
Desert sand	820
Glass (standard)	670
Argon	520
Iron and steel (average)	450
Zinc	387
Copper	385
Lead	129

Table.1 *Specific heat capacities*

Knowing the mass of a sample, its change in temperature and its specific heat capacity allows one to calculate the energy transferred.

$$\boxed{Q = mc\Delta T}$$

Where Q = heat transferred ($Joules$)
m = mass (kg)
c = specific heat capacity ($Jkg^{-1}K^{-1}$)
ΔT = $T_f - T_i$ = change in temperature ($K\ or$ °C)

Example 1
Find the amount of energy needed to raise the temperature of 3.0 kg of iron from 10°C to 18°C.

$Q = ?$
$m = 3.0\ kg$
$c = 450\ Jkg^{-1}K^{-1}$
$\Delta T = 18 - 10 = 8°\text{C}\ (or\ K)$

$$\begin{aligned} Q &= mc\Delta T \\ &= 3.0 \times 450 \times 8 \\ &= 10800\ J \\ &= 1.08 \times 10^4\ J \end{aligned}$$

Example 2
Find the amount of heat lost (to the surroundings) when 2.0 kg of water at 80°C cools to 20°C.

$Q = ?$
$m = 2.0\ kg$
$c = 4200\ Jkg^{-1}K^{-1}$
$\Delta T = 80 - 20 = 60°\text{C}\ (or\ K)$

$$\begin{aligned} Q &= mc\Delta T \\ &= 2.0 \times 4200 \times 60 \\ &= 504000\ J \\ &= 5.04 \times 10^5\ J \end{aligned}$$

Example 3
Find the increase in temperature when a 10.0 kg Aluminium bar is heated by 448500 $Joules$.

$\Delta T = ?$
$m = 10.0\ kg$
$c = 897\ Jkg^{-1}K^{-1}$
$Q = 448500\ J$

$$\begin{aligned} Q &= mc\Delta T \\ \therefore\ \Delta T &= \frac{Q}{mc} \\ &= \frac{448500}{10 \times 897} \\ &= 50°\text{C}\ (or\ K) \end{aligned}$$

Example 4
A student supplies $1.40 \times 10^4\ J$ of heat energy to 500 g of oil in a sealed container and the temperature rises from 10°C to 20°C.
Calculate the specific heat capacity (c) of the alcohol.

$c = ?$
$m = 0.5\ kg$
$Q = 1.40 \times 10^4\ J$
$\Delta T = 20 - 10 = 10°\text{C}\ (or\ K)$

$$\begin{aligned} Q &= mc\Delta T \\ \therefore\ c &= \frac{Q}{m\Delta T} \\ &= \frac{1.40 \times 10^4}{0.5 \times 10} \\ &= 2800\ Jkg^{-1}K^{-1} \end{aligned}$$

Calculating the Specific Heat Capacity

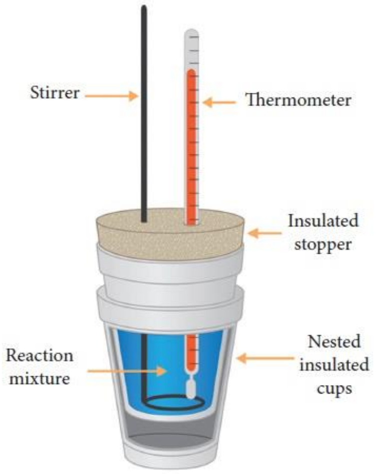


Figure.1 *Calorimeter*

To measure the specific heat capacity of a material we use a device called a **calorimeter**.

The steps are as follows:

1. measure the mass of the material being tested
2. measure the mass of the water in the calorimeter
3. measure the initial temperature of the material being tested
4. measure the initial temperature of the water in the calorimeter
5. place the material into the water filled calorimeter
6. stir the contents until thermal equilibrium is reached

Use the conservation of energy to calculate the specific heat capacity of the material being tested.

$$\boldsymbol{Energy\ lost\ by\ hotter\ =\ Energy\ gained\ by\ colder}$$
$$\boldsymbol{mc\Delta T\ (hotter) = mc\Delta T\ (colder)}$$

Example 5

A $75g$ block of aluminium at $100°C$ is placed in $200g$ of water at $10°C$, in a calorimeter. The final temperature of the mixture is $16°C$. Neglecting any heat energy gained by the calorimeter. What is the specific heat capacity of the aluminium?

Aluminium (hotter)	Water (colder)
$c = ?$	$c = 4200\ Jkg^{-1}K^{-1}$
$m = 0.075\ kg$	$m = 0.200\ kg$
$\Delta T = 100 - 16 = 84°C\ (or K)$	$\Delta T = 16 - 10 = 6°C\ (or K)$

NB: $\Delta T = T_{hot} - T_{cold}$

$$Energy\ lost\ by\ hotter\ =\ Energy\ gained\ by\ colder$$
$$mc\Delta T\ (hotter) = mc\Delta T\ (colder)$$

$$mc\Delta T\ (Al) = mc\Delta T\ (H_2O)$$
$$0.075 \times c \times 84 = 0.200 \times 4200 \times 6$$
$$\therefore c = \frac{0.200 \times 4200 \times 6}{0.075 \times 84}$$
$$= 800\ Jkg^{-1}K^{-1}$$

The specific heat capacity of the aluminium block was calculated to be $800\ Jkg^{-1}K^{-1}$.

Example 6

A hot iron bar of mass 2.5 kg, at a temperature 120°C, is placed into a full 10 litre (10 kg) container of water at 20°C. Find the final temperature of the water. What assumptions have you made?

<u>Iron (hotter)</u>
$c = 450\ Jkg^{-1}K^{-1}$
$m = 2.5\ kg$
$\Delta T = T_{hot} - T_{cold}$
$= 120 - T$

<u>Water (colder)</u>
$c = 4200\ Jkg^{-1}K^{-1}$
$m = 10.0\ kg$
$\Delta T = T_{hot} - T_{cold}$
$= T - 20$

NB: $T = T_{cold}(Fe) = T_{hot}(H_2O) = ?$

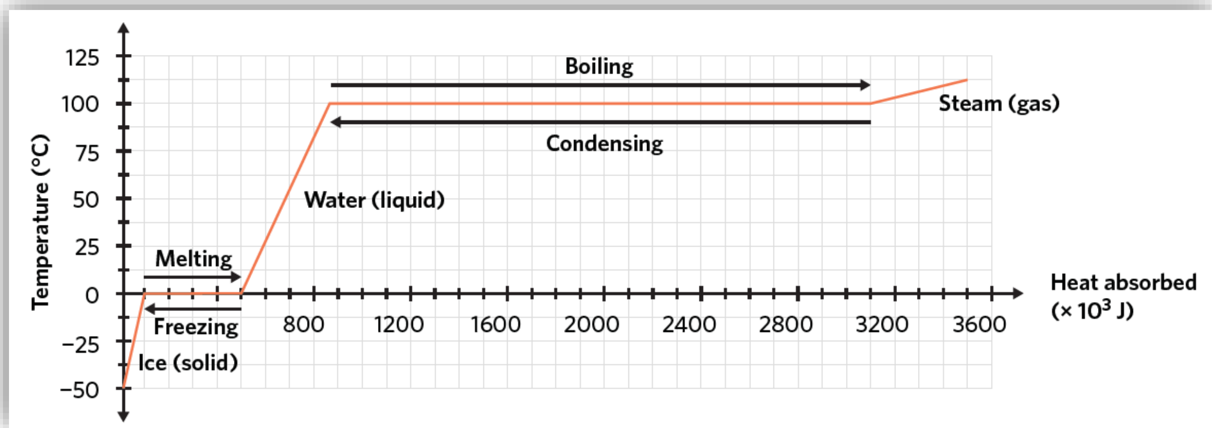
$Energy\ lost\ by\ hotter\ =\ Energy\ gained\ by\ colder$
$mc\Delta T\ (hotter) = mc\Delta T\ (colder)$

$mc\Delta T\ (Fe) = mc\Delta T\ (H_2O)$
$2.5 \times 450 \times (120 - T) = 10 \times 4200 \times (T - 20)$
$1125 \times (120 - T) = 42000 \times (T - 20)$
$(1125 \times 120) - (1125 \times T) = (42000 \times T) - (42000 \times 20)$
$135000 - 1125T = 42000T - 840000$
$135000 + 840000 = 42000T + 1125T$
$975000 = 43125T$
$\therefore T = \frac{975000}{43125}$
$= 22.6℃$

The equilibrium temperature of the iron bar and water is calculated to be 22.6℃.
This assumes that no heat is lost to the surroundings or transferred to the container.

Latent Heat

If you were to observe the temperature of a container of water as it approached boiling point, you would see it climb towards 100°C, reach 100°C at boiling point and go no further whilst all the water boiled away. Refer to **Graph.1** below.



Graph.1 *Heating curve for 1 kg of water*

Qn. How is it possible for heat energy to be added to a system and for the temperature of the water not to rise?

The answer is **latent heat**. The term "**latent**", literally means "**hidden**" and refers to the fact that adding heat energy does not change the temperature of the water while it is changing state. In other words it is heating without getting hot. This effect of adding heat without any change in temperature can be observed when water melts and boils.

The amount of **energy required** to **change 1 kg of a substance** from **solid to liquid, or liquid to solid**, at its melting point is called the **specific latent heat of fusion**.

The amount of **energy required** to change **1 kg of a substance** from **liquid to gas, or gas to liquid**, at its boiling point is called the **specific latent heat of vaporisation**.

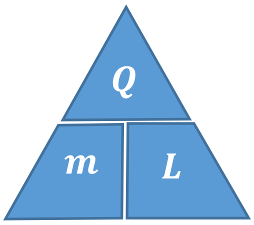
Substance	Specific latent heat of fusion (J kg^{-1})	Specific latent heat of vaporisation (J kg^{-1})
Water	3.3×10^5	2.3×10^6
Oxygen	6.9×10^3	1.1×10^5
Sodium chloride	4.9×10^5	2.9×10^6
Aluminium	2.2×10^3	1.7×10^4
Iron	2.8×10^5	6.3×10^6

Table.2 *Specific latent heat of fusion and vaporisation*

Knowing the mass of a sample and its specific latent heat capacity allows one to calculate the energy necessary to boil or melt a particular substance. The following equation is used:

$$\boxed{Q = mL}$$

Where Q = heat transferred (J)
m = mass (kg)
L = latent heat (Jkg^{-1})



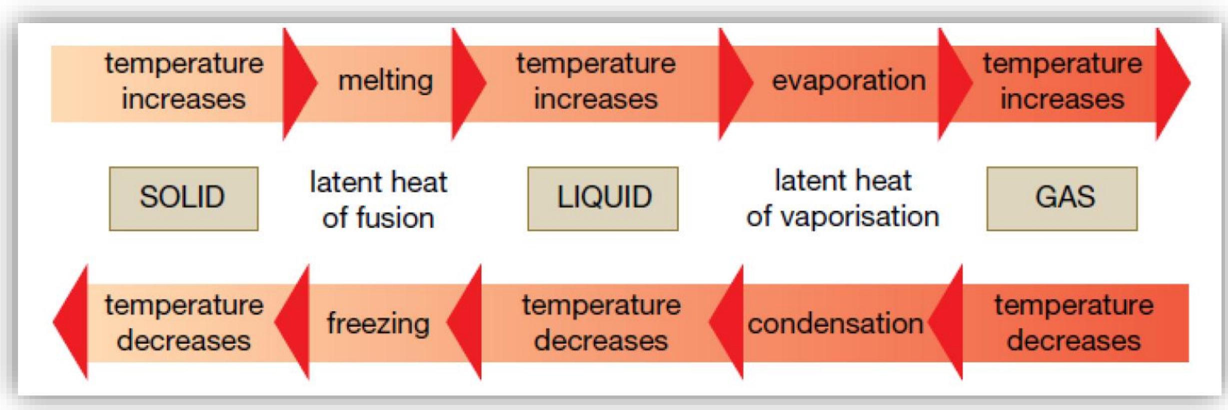


Figure.2 *Change of phase/state*

Example 7

Calculate the amount of heat energy necessary to change 1.5 kg of ice at 0°C to water at 0°C.
(As there is <u>no change in temperature</u>, this must be a latent heat question)

$c = ?$
$m = 1.5\ kg$
$L_f = 3.3 \times 10^5\ Jkg^{-1}$

$$\begin{aligned} Q &= mL_f \\ &= 1.5 \times 3.3 \times 10^5 \\ &= 495000\ J \\ &= 4.95 \times 10^5\ J \end{aligned}$$

Example 8

How much heat energy is needed to change 3.0 kg of water at 100°C completely into steam at 100°C.

$c = ?$
$m = 3.0\ kg$
$L_v = 2.3 \times 10^6\ Jkg^{-1}$

$$\begin{aligned} Q &= mL_v \\ &= 3.0 \times 2.3 \times 10^6 \\ &= 6900000\ J \\ &= 6.9 \times 10^6\ J \end{aligned}$$

Example 9
Calculate the amount of heat energy necessary to convert 2.5 kg of ice at −20°C to water at 40°C.

NB: This involves three steps:

1. The heating of ice from −20°C to 0°C
2. The change of phase from solid to liquid at 0°C
3. The heating of water from 0°C to 40°C

Heating of ice
(solid)
↑ temp

$Q = mc\Delta T$

$Q = ?$
$m = 2.5\ kg$
$c_{ice} = 2100\ Jkg^{-1}K^{-1}$
$\Delta T = 20°C$

$$\begin{aligned} Q &= mc\Delta T \\ &= 2.5 \times 2100 \times 20 \\ &= 105000\ J \\ &= 1.05 \times 10^5\ J \end{aligned}$$

Change of phase
(solid → liquid)
No Δ temp

$Q = mL_f$

$Q = ?$
$m = 2.5\ kg$
$L_f = 3.3 \times 10^5\ Jkg^{-1}$

$$\begin{aligned} Q &= mL_f \\ &= 2.5 \times 3.3 \times 10^5 \\ &= 825000\ J \\ &= 8.25 \times 10^5\ J \end{aligned}$$

Heating of water
(liquid)
↑ temp

$Q = mc\Delta T$

$Q = ?$
$m = 2.5\ kg$
$c_{water} = 4200\ Jkg^{-1}K^{-1}$
$\Delta T = 40°C$

$$\begin{aligned} Q &= mc\Delta T \\ &= 2.5 \times 4200 \times 40 \\ &= 420000\ \text{J} \\ &= 4.2 \times 10^5\ J \end{aligned}$$

∴ Total energy necessary $= 1.05 \times 10^5 + 8.25 \times 10^5 + 4.2 \times 10^5 = 1.35 \times 10^6\ J$

Evaporation
The molecules in a liquid are moving in a **random manner** and not all molecules are moving at the same speed. In fact the **speeds of the molecules vary considerably**.

The faster moving molecules near the surface of a liquid are able to **escape** from the surface of the liquid and become a gas, even though the temperature of the liquid is well below the boiling point. This is the process of **evaporation**.

This **reduces** the **average translational kinetic energy** of the remaining particles, and the liquid is **cooler** as a result. Refer to **Figure.3** below.

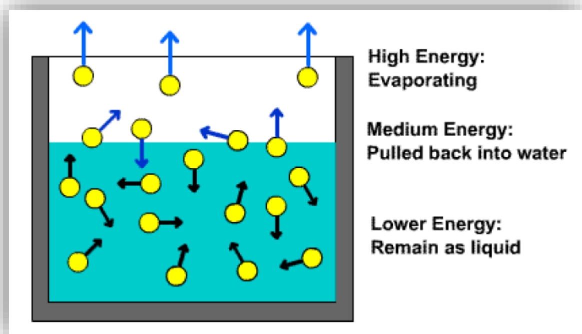


Figure.3 *Evaporation*

Exam Styled Questions

Question 1 (1 mark)

Considering the formula $c_{alcohol} = 2400\ Jkg^{-1}K^{-1}$, what is the change in temperature if $2184\ J$ of heat is gained by $70\ g$ of alcohol?

A. 168°C
B. 91°C
C. 13°C
D. 7°C

$\Delta Q = mc\Delta T$
2184 J = 0.07 × 2400 × ΔT
ΔT = 13°C

C

Question 2 (6 mark)

A Bunsen burner melts $5\ kg$ of ice originally found to have a temperature of $-15°C$. The melted water is then placed in a beaker and warmed to $15°C$. Assume negligible heat is contributed from the surroundings.

Data

H_2O (steam)	$c = 2000\ \text{J kg}^{-1}\ \text{K}^{-1}$
H_2O (liquid)	$c = 4200\ \text{J kg}^{-1}\ \text{K}^{-1}$
H_2O (ice)	$c = 2100\ \text{J kg}^{-1}\ \text{K}^{-1}$
H_2O	$l_f = 3.34 \times 10^5\ \text{J kg}^{-1}$
H_2O	$l_v = 2.25 \times 10^6\ \text{J kg}^{-1}$

What is the minimum amount of energy delivered by the Bunsen burner?

$\Delta Q = mc\Delta T$
$= 5 \times 2100(0 - -15)$ 1 mark
$= 157\ 500$ J 1 mark

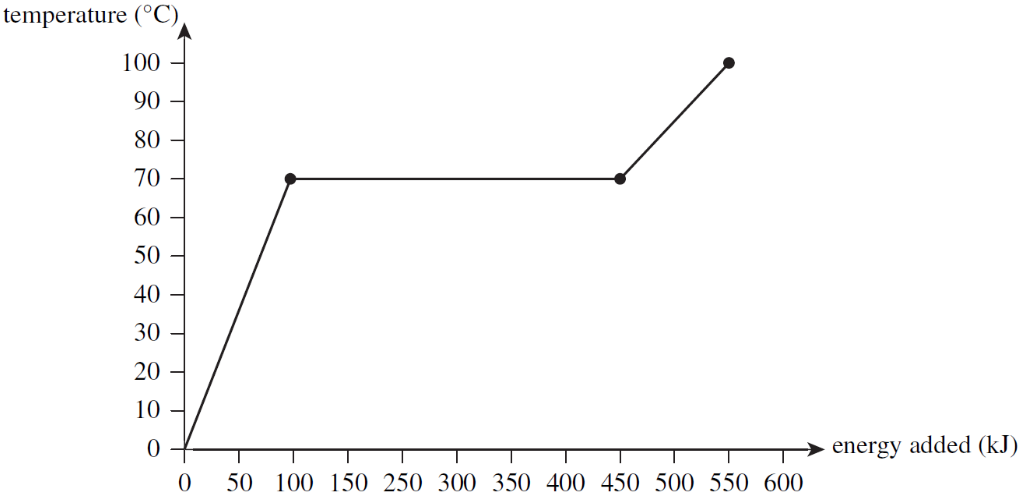
$\Delta Q = mL_{\text{fusion}}$
$= 5 \times 3.34 \times 10^5$ 1 mark
$= 1\ 670\ 000$ J 1 mark

$\Delta Q = mc\Delta T$
$= 5 \times 4200(15 - 0)$ 1 mark
$= 315\ 000$ J

$\Delta Q = 157\ 500 + 1\ 670\ 000 + 315\ 000$
$= 2\ 142\ 500$ J
$= 2.1 \times 10^6$ J 1 mark

The following information is relevant for Questions 3 to 6

The below figure shows the heating curve for a $2.0\ kg$ sample of material that begins as a solid at room temperature and finishes as a hot liquid. Note that energy is added to the material at a constant rate.



Question 3 (1 mark)
What is the melting point of the material?

$\boxed{70\ ℃}$

Question 4 (2 marks)
Why does the substance's temperature remain at $70°C$ for a period of time, even though $350\ kJ$ is added at a constant rate?

The substance is changing from a solid to liquid. 1 mark

There is no change in the temperature and the energy increases the potential energy of the particles in the material by reducing the inter-particle/intermolecular forces. 1 mark

Question 5 (2 marks)
Calculate the latent heat of fusion for the sample.

$$\Delta Q = mL_{\text{fusion}}$$

$$L_{\text{fusion}} = \frac{\Delta Q}{m}$$

$$= \frac{350\ 000}{2} \quad \text{1 mark}$$

$$= 175\ 000\ \text{J kg}^{-1}$$

$$= 1.8 \times 10^5\ \text{J kg}^{-1} \quad \text{1 mark}$$

$$\boxed{1.8 \times 10^5 Jkg^{-1}}$$

Question 6 (2 marks)

Calculate the specific heat capacity of the sample while it is a liquid.

$$\Delta Q = mc\Delta T$$

$$c = \frac{\Delta Q}{m\Delta T}$$ 1 mark

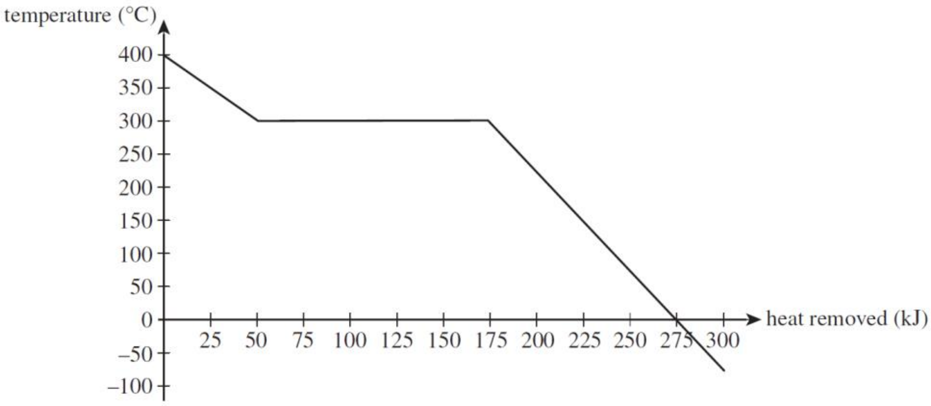
$$= \frac{550\,000 - 450\,000}{2 \times 30}$$

$$= 1.7 \times 10^{3}\ \text{J kg}^{-1}\ {}^{\circ}\text{C}^{-1}$$ 1 mark

$$\boxed{1.7 \times 10^{3} Jkg^{-1}K^{-1}}$$

The following information is relevant for Questions 7 to 9

The below figure shows a temperature ($°C$) versus heat removed (kJ) graph for a 500 g substance that begins an experiment as a liquid.



Question 7 (1 mark)
What is the freezing point of the substance?

$\boxed{300 \ ℃}$

Question 8 (2 marks)
Calculate the latent heat of fusion for the sample.

$$\Delta Q = mL_{\text{fusion}}$$

$$L_{\text{fusion}} = \frac{\Delta Q}{m}$$

$$= \frac{125\ 000}{0.5}$$ 1 mark

$$= 250\ 000 \text{ J kg}^{-1}$$

$$= 2.5 \times 10^{2} \text{ kJ kg}^{-1}$$ 1 mark

$\boxed{\mathbf{2.5 \times 10^{2} kJkg^{-1}}}$

Question 9 (2 marks)
Calculate the specific heat capacity of the sample while it is a solid.

$$\Delta Q = mc\Delta T$$

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

$$= \frac{100\ 000}{0.5 \times 300}$$ 1 mark

$$= 666.67 \text{ J kg}^{-1}\ °\text{C}^{-1}$$

$$= 6.7 \times 10^{2} \text{ kJ kg}^{-1}\ °\text{C}^{-1}$$ 1 mark

$\boxed{\mathbf{6.7 \times 10^{2} kJkg^{-1}K^{-1}}}$

Question 10 (5 marks)
A Bunsen burner is used to completely melt a 200 g sample of paraffin wax. The paraffin wax is initially at room temperature ($25°C$).
What is the minimum amount of energy delivered by the Bunsen burner if the melting point of the candle is $46°C$? (Assume negligible heat is contributed from the surroundings.)

Note: $c = 2.2 \times 10^3\ Jkg^{-1}K^{-1}$ and $L_f = 2.0 \times 10^5\ Jkg^{-1}$

$\Delta Q = mc\Delta T$

$= 0.2 \times 2.2 \times 10^3 (46 - 25)$ 1 mark

$= 9240$ J 1 mark

$\Delta Q = mL_{\text{fusion}}$

$= 0.2 \times 2.0 \times 10^5$ 1 mark

$= 40\ 000$ J 1 mark

$\Delta Q = 9240 + 40\ 000$

$= 49\ 240$ J

$= 4.9 \times 10^5$ J 1 mark

$$\boxed{4.9 \times 10^5\ J}$$

Question 11 (1 mark)
In an experiment, a 1.0 kg aluminium block is heated to $90°C$. It is then dropped into 5.0 kg of water at $20°C$. The specific heat capacity of water is 4200 $Jkg^{-1}K^{-1}$ and the specific heat capacity of aluminium is 880 $Jkg^{-1}K^{-1}$.

Assuming no energy is transferred to the surrounding air or container, the final temperature of the aluminium block and water is closest to

A. 21°C
B. 23°C
C. 35°C
D. 55°C

$\boxed{\text{B}}$

$$\Delta Q_{\text{water}} = \Delta Q_{\text{aluminium}}$$
$$5.0 \times 4200\,(T - 20°C) = 1.0 \times 880(90 - T)$$
$$21\ 000T - 420\ 000 = 79\ 200 - 880T$$
$$21\ 880T = 499\ 200$$
$$T = 22.82°C$$
$$T = 23°C$$

Question 12 (1 mark)

$4.00\ g$ of silver in liquid form solidifies at a constant temperature. The latent heat of fusion of silver is $1.05 \times 10^5\ Jkg^{-1}$.

How much energy is removed when this change is made?

A. $420\ J$
B. $2.60 \times 10^4\ J$
C. $4.20 \times 10^5\ J$
D. $2.60 \times 10^7\ J$

$$Q = ml_f$$
$$= 0.00400 \times 1.05 \times 10^5$$
$$= 420 \text{ J}$$

A

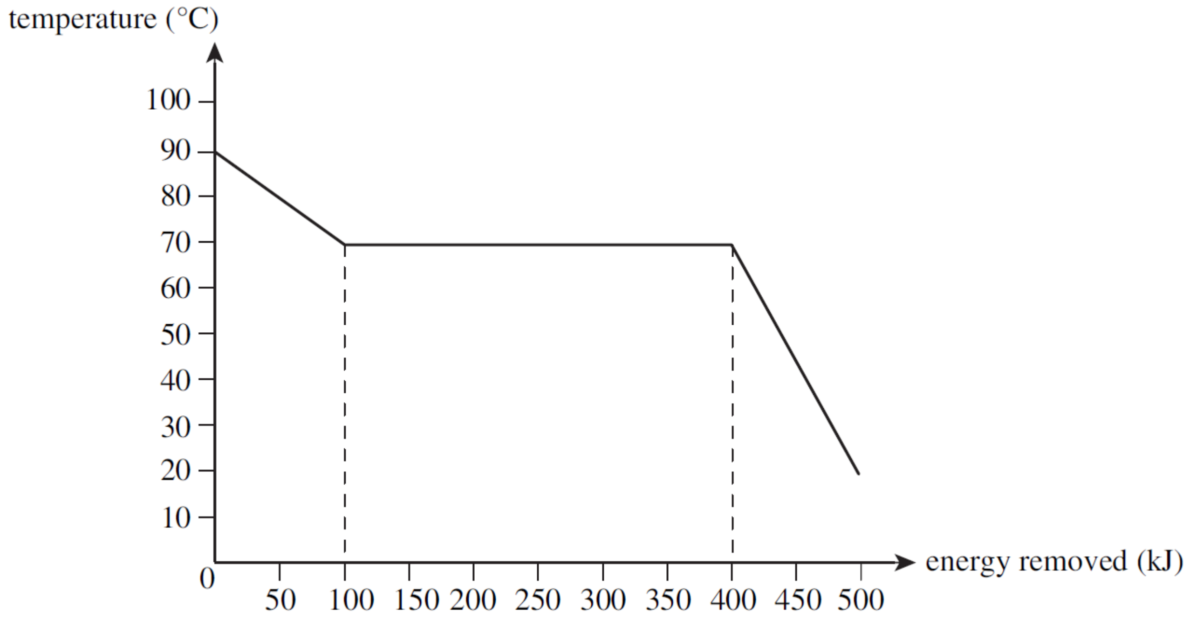
Question 13 (2 mark)

Explain why the evaporation of sweat results in the cooling of the body.

The evaporation of sweat causes sweat to change states from a liquid to a vapour. 1 mark
This takes away energy from the body as latent heat, giving a cooling effect. 1 mark

The following information is relevant for Questions 14 to 16

The below figure shows a temperature ($°C$) versus energy removed (kJ) graph for 200 g of a substance that begins an experiment as a liquid and finishes as a solid. Energy is removed from the material at a constant rate.



Question 14 (1 mark)
What happens to the substance when it reaches 70°C?

The substance changes state from a liquid to a solid. 1 mark

Question 15 (2 mark)
Calculate the latent heat of fusion of the substance.

$$Q = ml_{\text{fusion}}$$

$$300\,000 = 0.20 \times l_{\text{fusion}}$$ 1 mark

$$l_{\text{fusion}} = 1\,500\,000 \text{ J kg}^{-1}$$ 1 mark

$$\boxed{\mathbf{1.5 \times 10^6 Jkg^{-1}}}$$

Question 16 (2 mark)
Calculate the specific heat capacity of the substance when it is a liquid.

$$Q = mc\Delta T$$

$$100\,000 = 0.2 \times c \times (90 - 70)$$ 1 mark

$$c = 25000 \text{ J kg}^{-1}\text{K}^{-1}$$ 1 mark

$$\boxed{\mathbf{2.5 \times 10^4 Jkg^{-1}K^{-1}}}$$