Physics Year 11 Week 9

SHC and Latent Heat

Background

Heat capacity, ***C***, of a body is defined as the amount of heat (Q) required to raise its temperature (θ) by one degree Kelvin (or 1 °C), without going through a change in state.

SI unit of *heat capacity* is *joule per kelvin* or *joule per degree Celsius* [$\mathbf{J\,K^{-1}}$ or $\mathbf{J\,°C^{-1}}$].

$$C = \frac{Q}{\Delta T}$$

C = heat capacity ($\mathrm{J\,K^{-1}}$, $\mathrm{J\,°C^{-1}}$)

Q = heat or thermal energy absorbed or released (J)

ΔT = change in temperature (°C or K)

$$\Delta T = \left(T_{\text{final}} - T_{\text{initial}}\right)$$

Specific heat capacity

Specific heat capacity, *c*, of a body is defined as the amount of heat (Q) required to raise the temperature (θ) of a unit mass of it by one degree, without going through a change in state.

$$\begin{bmatrix}\text{thermal}\\ \text{energy}\end{bmatrix} = [\text{mass}] \times \begin{bmatrix}\text{specific}\\ \text{heat}\\ \text{capacity}\end{bmatrix} \times \begin{bmatrix}\text{change in}\\ \text{temperature}\end{bmatrix}$$

$$c = \frac{C}{m}$$

$$Q = m \times c \times \Delta T = \left(T_{\text{final}} - T_{\text{initial}}\right)$$

c = specific heat capacity ($\text{J kg}^{-1}\,\text{K}^{-1}$, $\text{J kg}^{-1}\,{}^{\circ}\text{C}^{-1}$)

C = heat capacity (J K^{-1} or $\text{J}\,{}^{\circ}\text{C}^{-1}$)

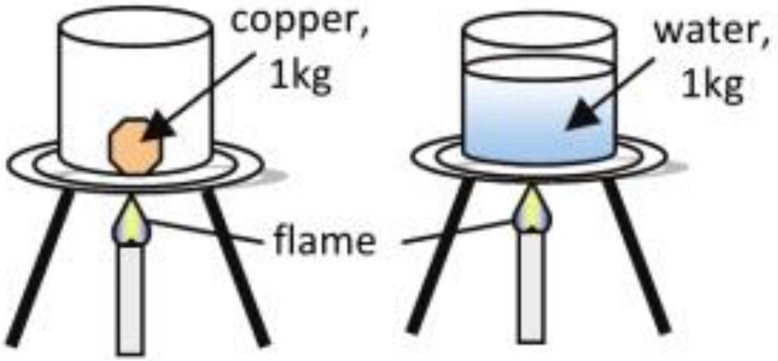
m = mass of the substance (kg)

Q = heat or thermal energy absorbed or released (J)

Differences

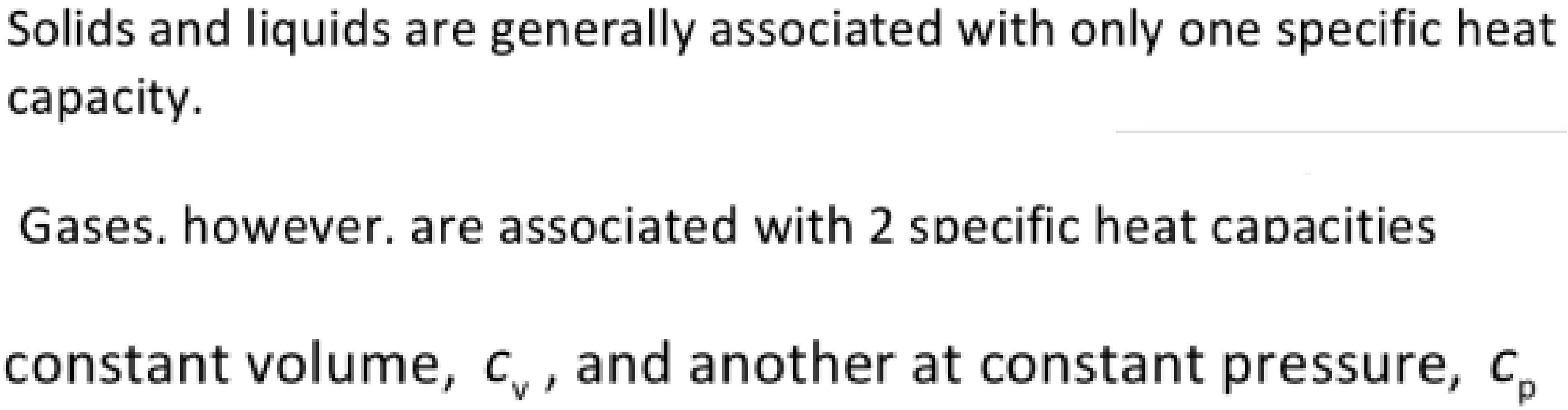
Specific heat capacity of gases is higher than that of liquids and much higher than that of solids. The substances with higher *specific heat capacity* cool or warm very slowly compared to the substances with lower *specific heat capacity*.

The *specific heat capacity* of water (liquid) is $4200\ \text{J kg}^{-1}\,{}^{\circ}\text{C}^{-1}$ and for copper (solid) it is $400\ \text{J kg}^{-1}\,{}^{\circ}\text{C}^{-1}$.



copper takes less time than water to increase its temperature by say, 5 °C

Differences continued



Solids and liquids are generally associated with only one specific heat capacity.

Gases, however, are associated with 2 specific heat capacities

constant volume, c_v , and another at constant pressure, c_p

Water

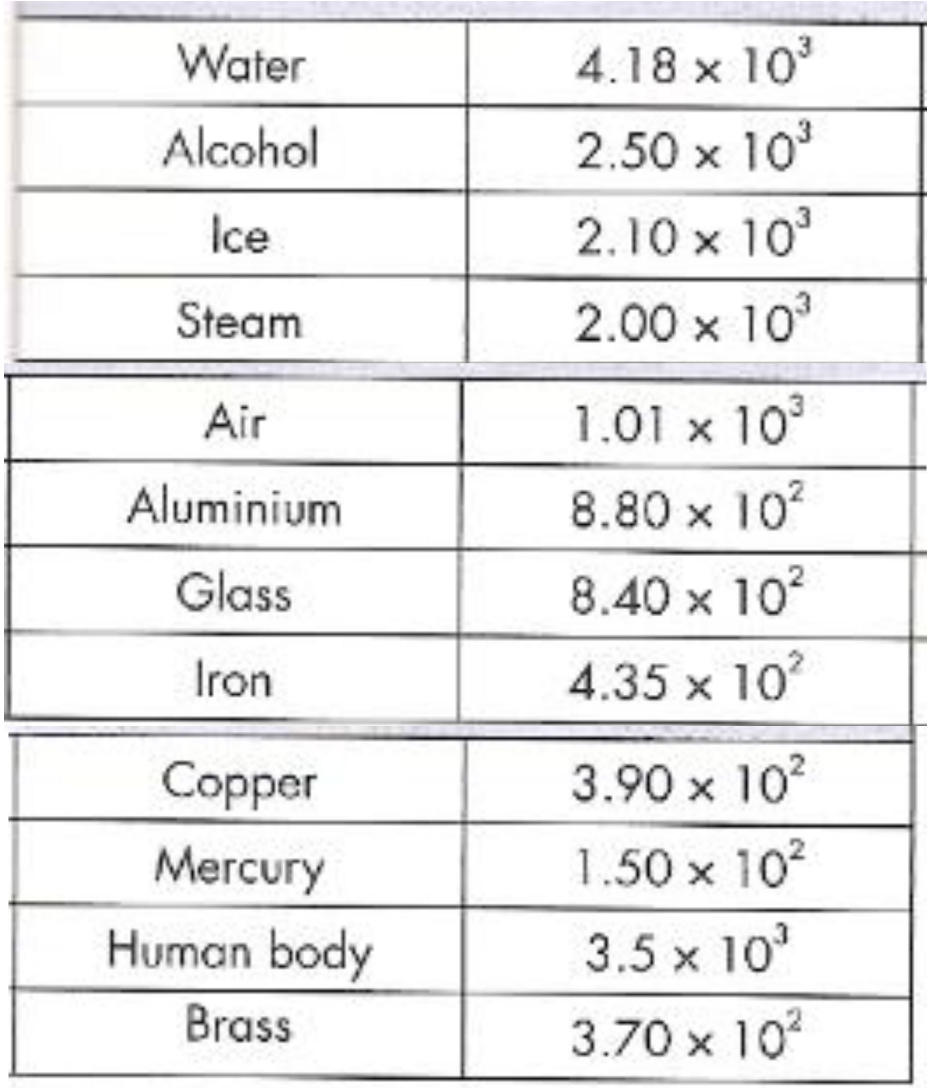
Water is suitable for use as **coolant** on cars.

The s.h.c. of the seas is higher than the s.h.c. of the land masses, it leads to **milder climate** in the coastal areas, *i.e.*, a favourite settlement choice for man, plants and animals alike.

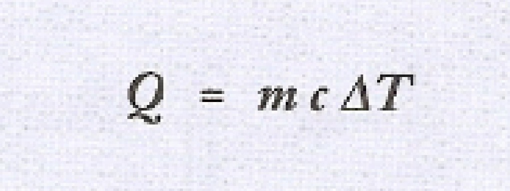
Warm-blooded animals are possible as the large percentage of water in the body keeps the animal’s temperature stable with little fluctuations.

Water is used as **fire extinguisher** due to its high heat capacity.

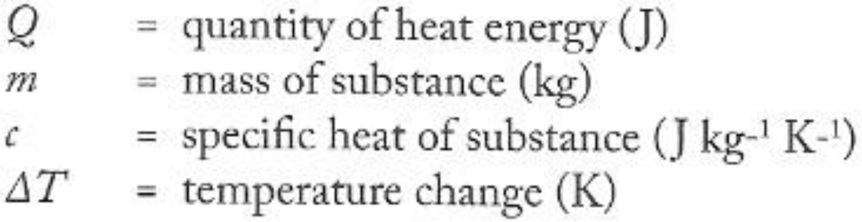
SHC of Common Substances



Water	4.18×10^3
Alcohol	2.50×10^3
Ice	2.10×10^3
Steam	2.00×10^3
Air	1.01×10^3
Aluminium	8.80×10^2
Glass	8.40×10^2
Iron	4.35×10^2
Copper	3.90×10^2
Mercury	1.50×10^2
Human body	3.5×10^3
Brass	3.70×10^2



$$Q = mc\Delta T$$



Q = quantity of heat energy (J)
m = mass of substance (kg)
c = specific heat of substance (J kg^{-1} K^{-1})
ΔT = temperature change (K)

Worked Example 1

A kettle is filled with tap water which is at a temperature of 22.0°C. If the mass of the water is 1.85 kg calculate the heat required to raise its temperature to 100.0°C.

Q = ?

m = 1.85 kg

c = 4180 J kg^{-1} K^{-1}

ΔT = (100 – 22.0) °C

$$
\begin{aligned}
Q &= m\,c\,\Delta T \\
&= (1.85)(4180)(100 - 22.0) \\
&= 6.03 \times 10^{5}\ \text{J} \\
&= 6.03 \times 10^{2}\ \text{kJ}
\end{aligned}
$$

Worked Example 2

A home heater supplies 1.80 MJ of heat in 15.0 minutes to a room which contains 95.5 kg of air. Assuming that all of the heat was absorbed by the air only, find the rise in temperature in the room.

What is the power rating of the heater?

$Q = 1.80 \times 10^6$ J

$m = 95.5$ kg

$c = 1.01 \times 10^3$ J kg^{-1} K^{-1}

$\Delta T = ?$

$t = 15.0$ min = 900 s

$P = ?$

$$Q = m\,c\,\Delta T$$

$$\therefore \Delta T = \frac{Q}{mc}$$

$$\Delta T = \frac{1.80 \times 10^6}{(95.5)(1010)}$$

$$= 18.7^\circ\mathrm{C}$$

Worked Example 2 continued

$$\text{Power} = \frac{\text{energy}}{\text{time}}$$

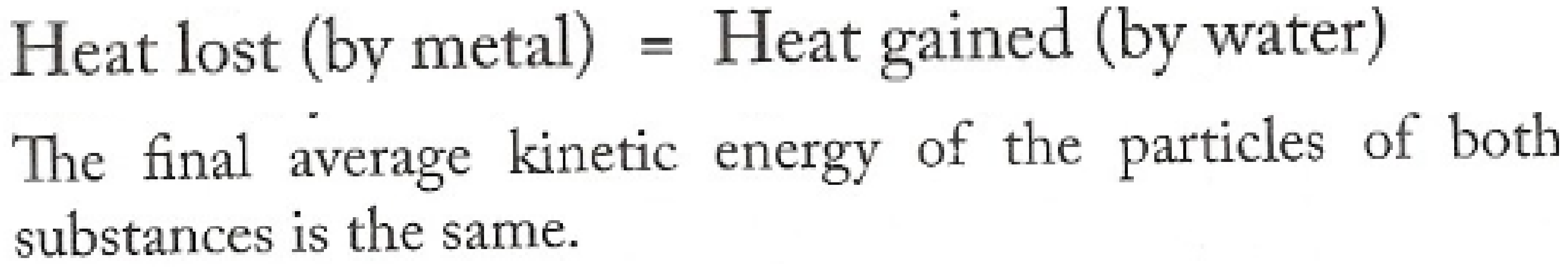
$$= \frac{1.80 \times 10^6}{900}$$

$$= 2000\ \text{Js}^{-1}$$

$$= 2.00\ \text{kW}$$

Thermal Equilibrium

The principle of conservation of energy states that energy can be neither created nor destroyed, but can only be converted from one form to another.
If there is no heat loss to the external surroundings, the heat lost by the hotter objects equals the heat gained by the cooler objects when placed in thermal contact.

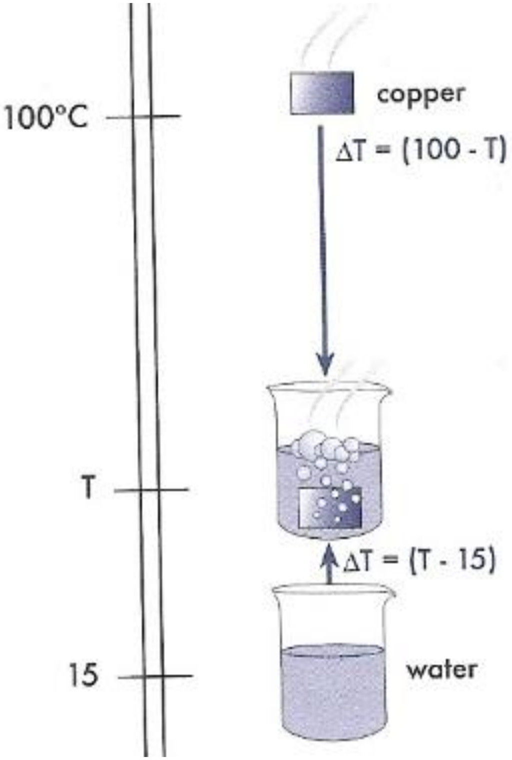


Heat lost (by metal) = Heat gained (by water)

The final average kinetic energy of the particles of both substances is the same.

Worked Example 3

A 650 g block of copper metal is initially heated to 100°C (in boiling water), and then placed into 500 g of cold water at 15.0°C.



For copper:

$m = 650\ \text{g}$

$c = 390\ \text{J kg}^{-1}\ \text{K}^{-1}$

$\Delta T = 100 - T$

For water:

$m = 500\ \text{g}$

$c = 4180\ \text{J kg}^{-1}\ \text{K}^{-1}$

$\Delta T = T - 15.0\ ^{\circ}\text{C}$

Worked Example 3 continued

Heat lost by copper = $m\,c\,\Delta T$ = $(0.650)(390)(100 - T)$

Heat gained by water = $m\,c\,\Delta T$ = $(0.500)(4180)(T - 15)$

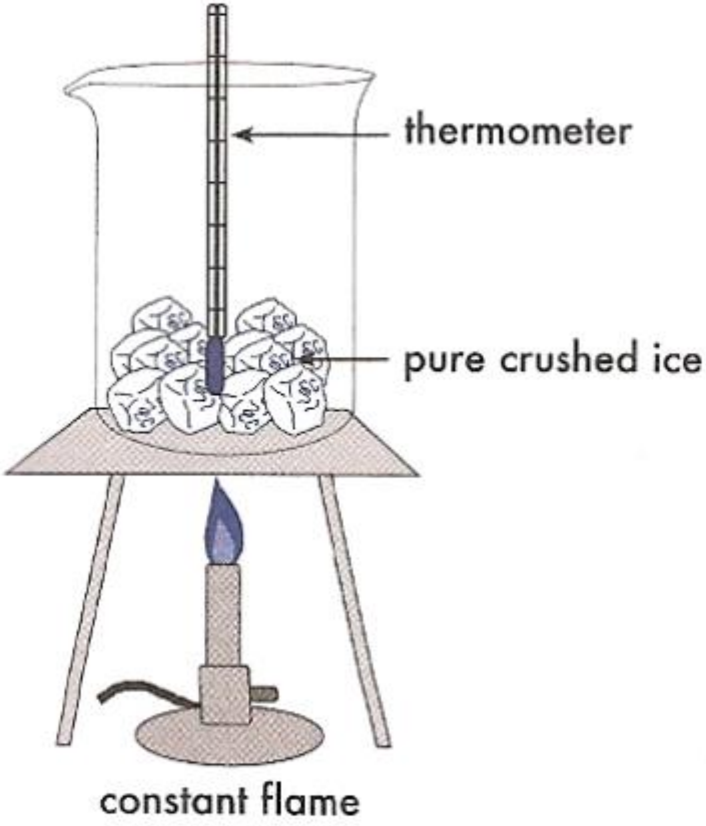
Heat lost by copper = $25350 - 253.5T$

Heat gained by water = $2090T - 31350$

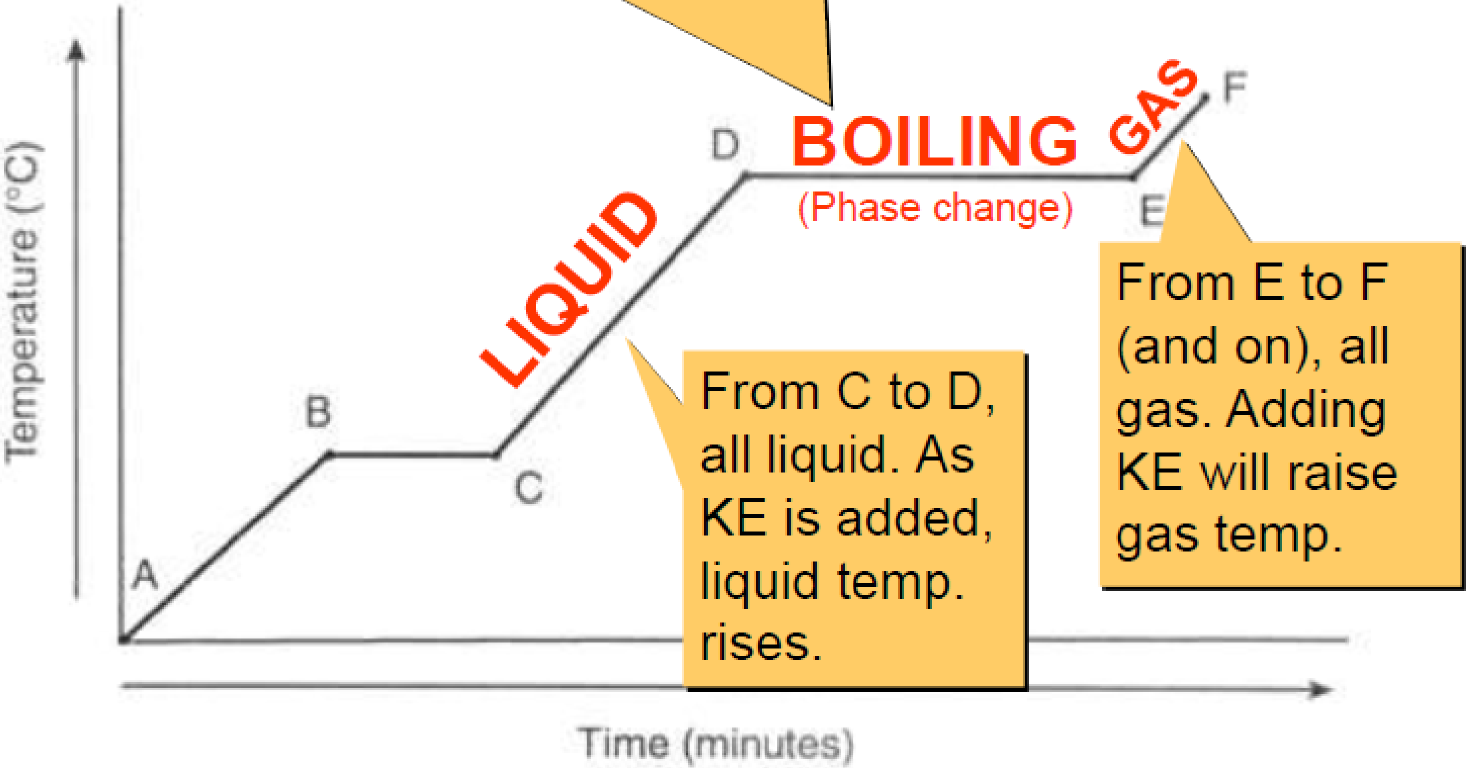
Since Heat lost = Heat gained

$$\begin{aligned} 25350 - 253.5T &= 2090T - 31350 \\ 2343.5T &= 56700 \\ T &= 24.2\ ^\circ\mathrm{C} \end{aligned}$$

Change of State



Kinetic Energy & Changes of State Liquid/Gas



When liquid reaches BP (D), all energy goes into separating particles from each other. Temp stays flat.

From C to D, all liquid. As KE is added, liquid temp. rises.

From E to F (and on), all gas. Adding KE will raise gas temp.

Specific Latent Heat

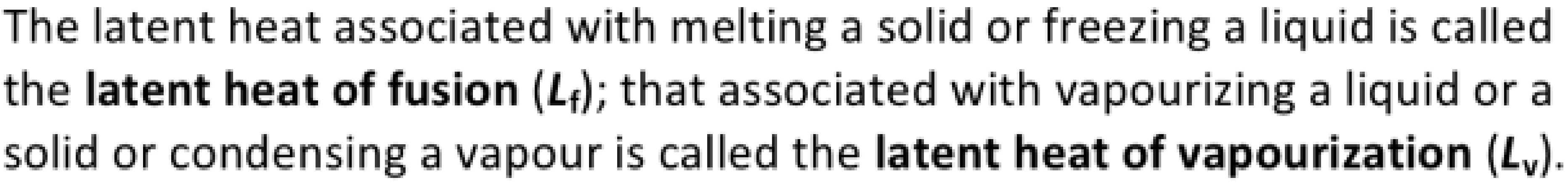
Latent heat of a substance is the amount of energy absorbed or released by the substance during a change in its *physical state* that occurs without changing its *temperature*.

$$L = Q \qquad \Rightarrow \quad Q = L$$

Q = amount of thermal energy (heat) absorbed or released (joule)

L = latent heat of the body (joule)

LHF and LHV



The latent heat associated with melting a solid or freezing a liquid is called the **latent heat of fusion** (L_f); that associated with vapourizing a liquid or a solid or condensing a vapour is called the **latent heat of vapourization** (L_v).

Specific latent heat of fusion

Specific latent heat of fusion, l_f, of a substance is defined as the amount of heat required to change unit mass of the substance from solid to liquid state, without changing its *temperature*.

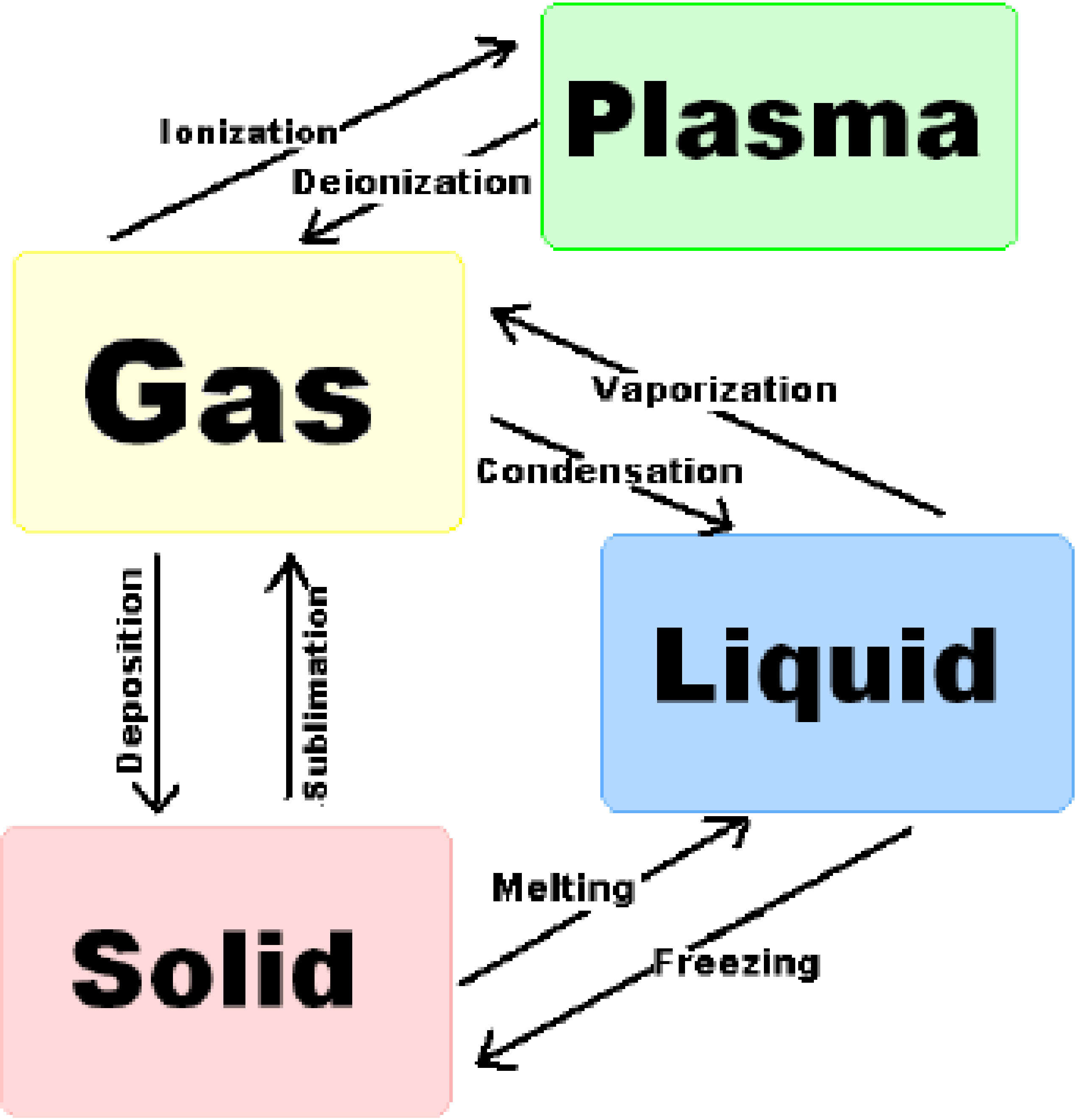
$$l_f = \frac{Q}{m} \Rightarrow Q = m \times l_f = L_f$$

Q = amount of thermal energy (heat) absorbed or released

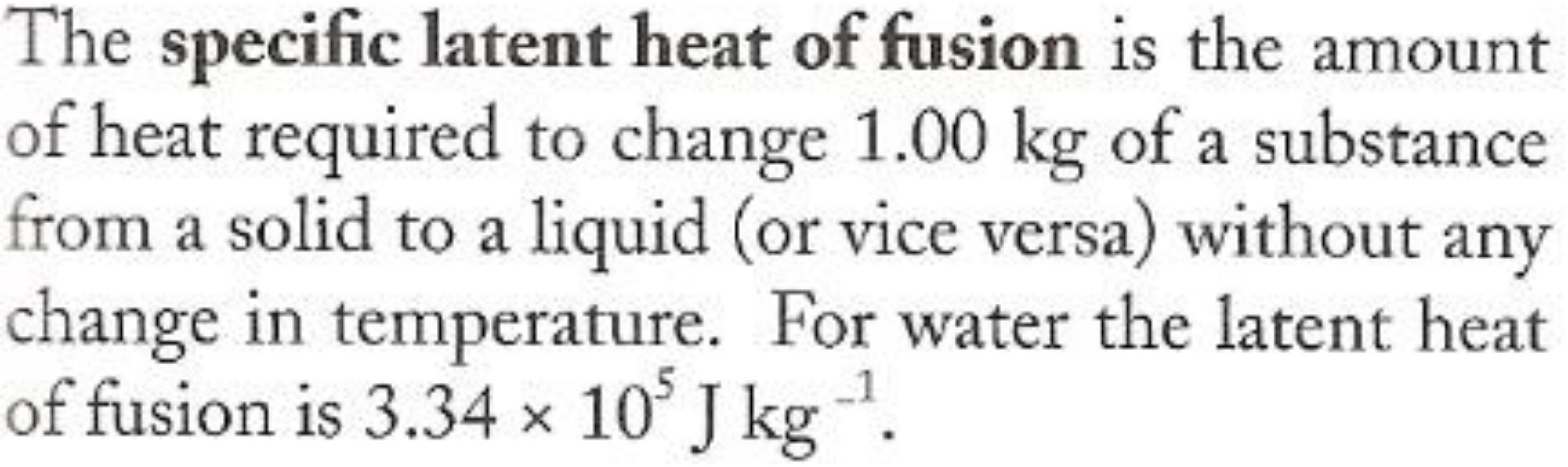
m = mass of the substance

l_f = specific latent heat of fusion

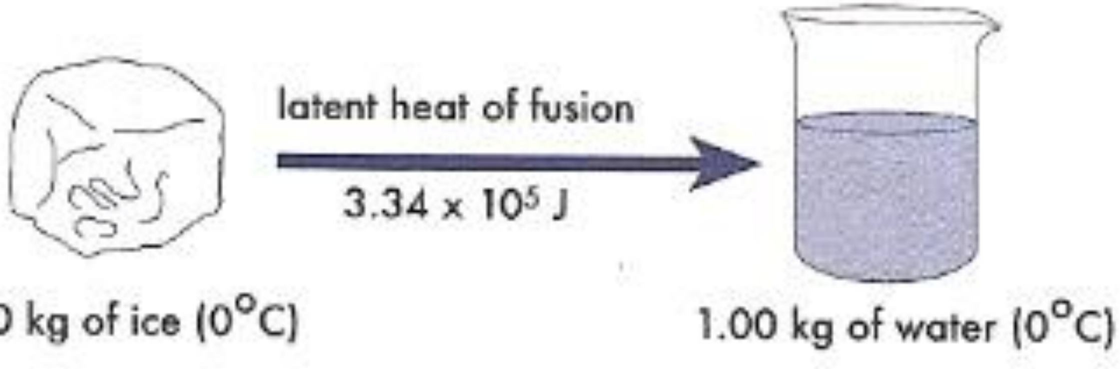
L_f = latent heat of fusion


Plasma
Ionization
Deionization
Gas
Vaporization
Condensation
Liquid
Deposition
Sublimation
Solid
Melting
Freezing

Latent Heat of Fusion

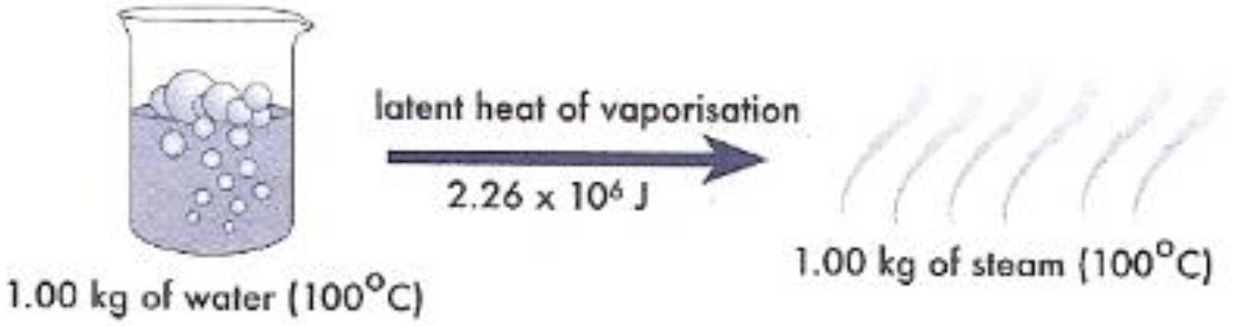


The **specific latent heat of fusion** is the amount of heat required to change 1.00 kg of a substance from a solid to a liquid (or vice versa) without any change in temperature. For water the latent heat of fusion is 3.34×10^5 J kg^{-1}.



Latent Heat of Vaporisation

The **specific latent heat of vaporisation** is the amount of heat required to change 1.00 kg of a substance from a liquid to a gas (or vice versa) without any change in temperature. For water the latent heat of vaporisation is $2.26\times 10^{6}\ \mathrm{J\ kg^{-1}}$.



LH of S

- The latent heat of sublimation is the heat needed to change each kilogram of solid to gas, or gas to solid.

Specific Latent Heat

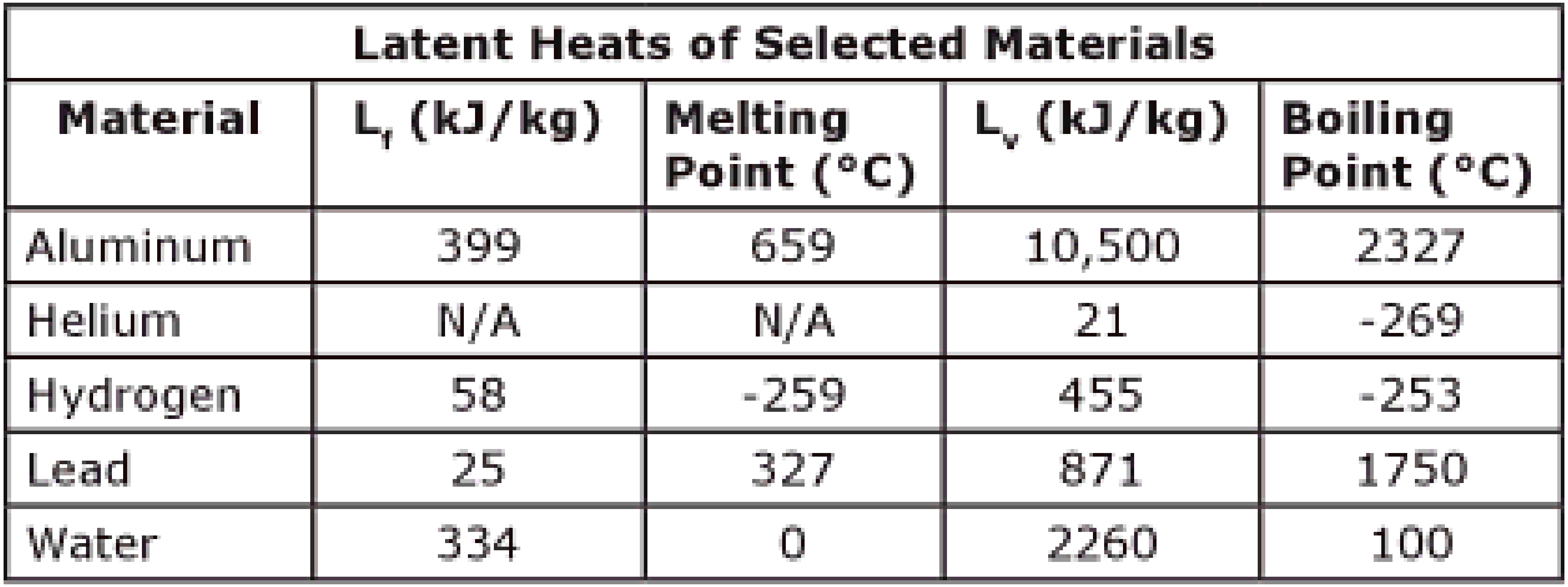
$$Q = mL$$

Q = quantity of heat required or given off (J)

m = mass of substance (kg)

L = specific latent heat (J kg^{-1})

Latent Heats of Selected Materials

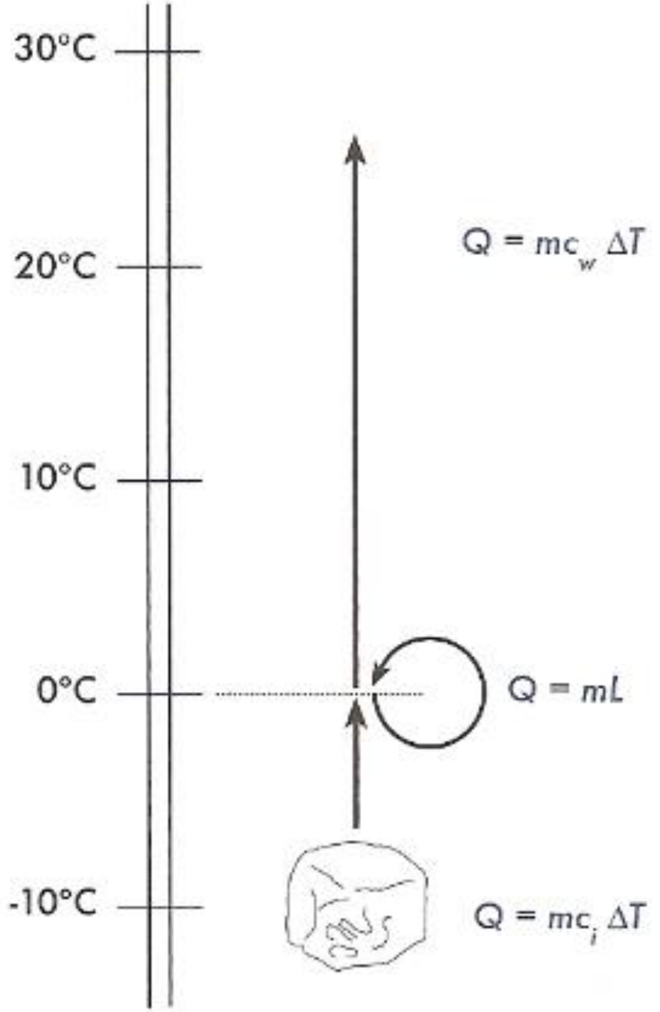


Latent Heats of Selected Materials

Material	L_f (kJ/kg)	Melting Point (°C)	L_v (kJ/kg)	Boiling Point (°C)
Aluminum	399	659	10,500	2327
Helium	N/A	N/A	21	-269
Hydrogen	58	-259	455	-253
Lead	25	327	871	1750
Water	334	0	2260	100

Worked Example 4

Determine the amount of heat required to change 250 g of ice at –10.0°C to water at 25.0°C.



Worked Example 4 continued

- Heat required to change the temperature of the ice from −10.0°C to 0°C.

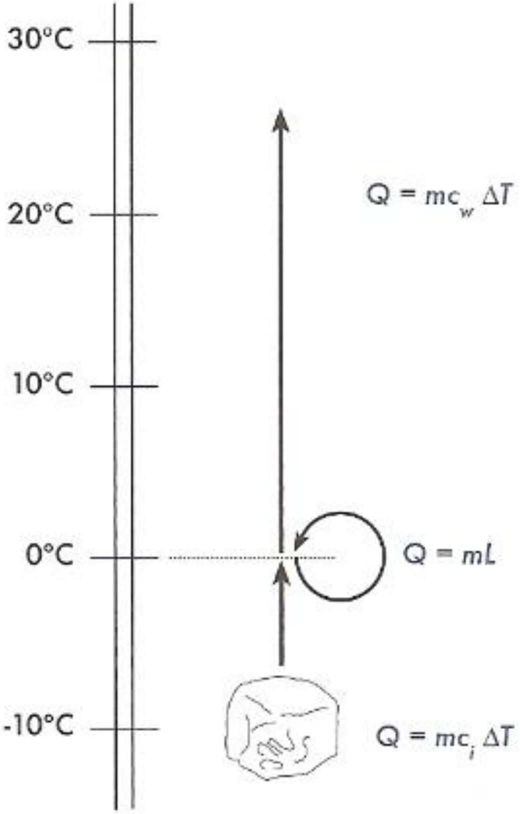
$$\begin{aligned} Q_1 &= m\, c_i \Delta T \\ &= (0.250)\,(2100)\,(10) \\ &= 5.25 \times 10^3 \text{ J} \end{aligned}$$

- Heat required to change the ice at 0°C to water at 0°C.

$$\begin{aligned} Q_2 &= m\, L \\ &= (0.250)\,(3.34 \times 10^5) \\ &= 8.35 \times 10^4 \text{ J} \end{aligned}$$

- Heat required to change the temperature of the water from 0°C to 25.0°C.

$$\begin{aligned} Q_3 &= m\, c_w \Delta T \\ &= (0.250)\,(4180)\,(25.0) \\ &= 2.61 \times 10^4 \text{ J} \end{aligned}$$



Hence total heat energy required:

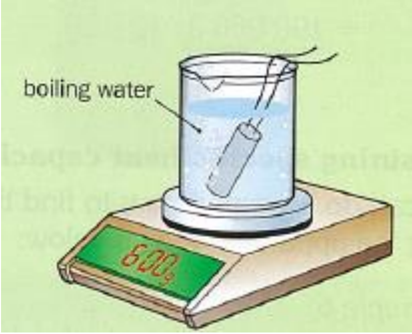
$$\begin{aligned} Q_T &= Q_1 + Q_2 + Q_3 \\ &= 1.15 \times 10^5 \text{ J} \end{aligned}$$

Worked Example 5

The power of the immersion heater in the diagram is 60 W.
In 5 minutes, the top pan balance reading falls from 282 g to 274 g.
What is the specific latent heat of vaporisation of water?

Energy, ΔQ = power × time = 60 W × (5 × 60) s = 18 000 J

Mass of water evaporated = 282 g – 274 g = 8 g = 8×10^{-3} kg

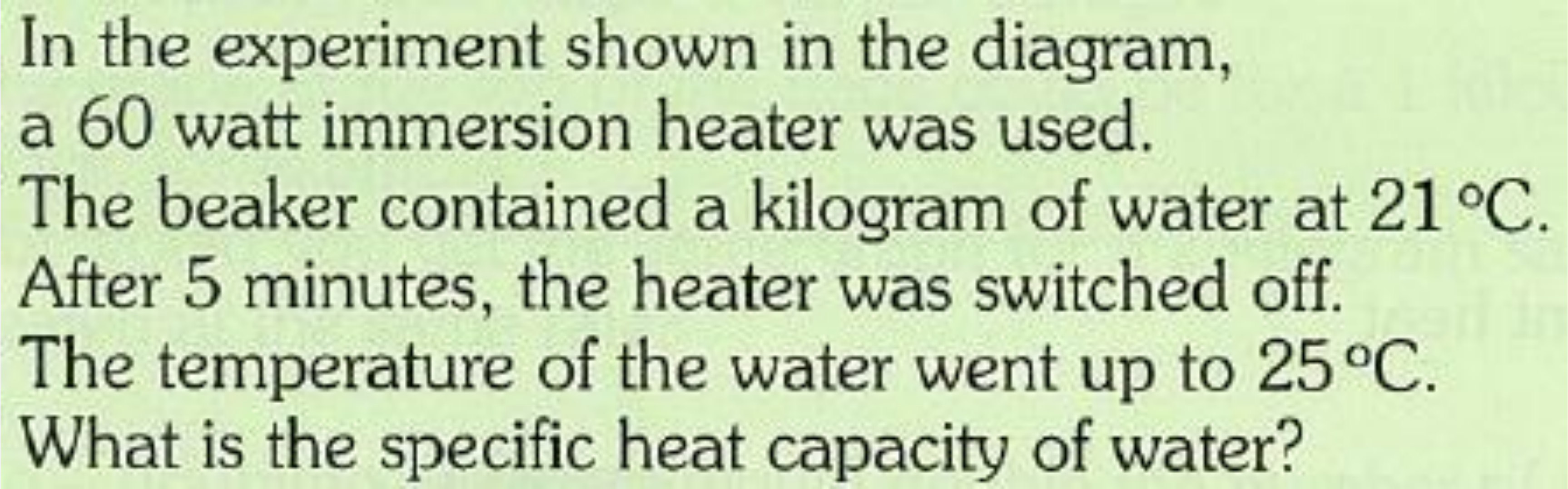


$$\Delta Q = mL$$

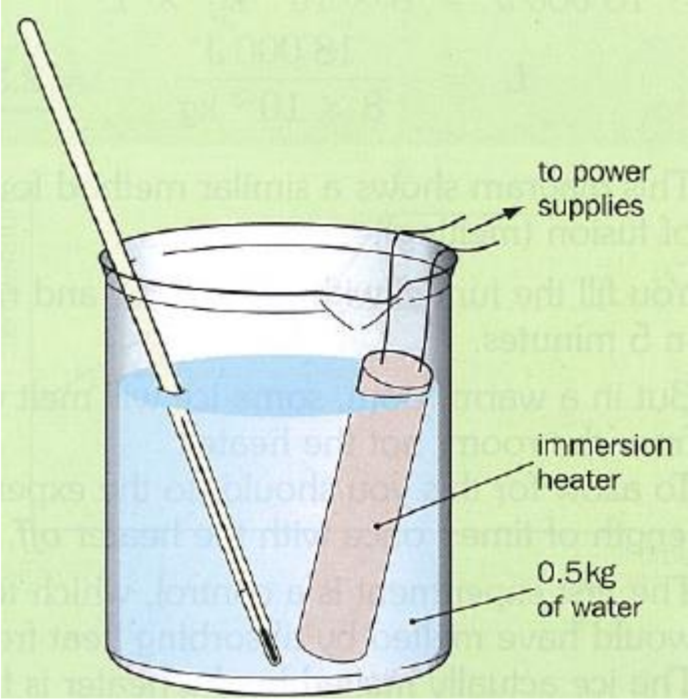
$$18\,000 \text{ J} = 8 \times 10^{-3} \text{ kg} \times L$$

$$L = \frac{18\,000 \text{ J}}{8 \times 10^{-3} \text{ kg}} = \underline{2.3 \times 10^{6} \text{ J kg}^{-1}}$$

Worked Example 6



In the experiment shown in the diagram,
a 60 watt immersion heater was used.
The beaker contained a kilogram of water at 21 °C.
After 5 minutes, the heater was switched off.
The temperature of the water went up to 25 °C.
What is the specific heat capacity of water?



Worked Example 6 continued

Change in internal energy, $= \text{power} \times \text{time}$

$= 60\ \text{W} \times (5 \times 60)\ \text{s} = 18\,000\ \text{J}$

The rise in temperature $= 25\,^{\circ}\text{C} - 21\,^{\circ}\text{C} = 4\,^{\circ}\text{C} = 4\ \text{K}$

$$= \boldsymbol{m\ c\ \Delta T}$$

$18\,000\ \text{J} = 1\ \text{kg} \times c \times 4\ \text{K}$

$$c = \frac{18\,000\ \text{J}}{4\ \text{kg K}} = \underline{4500\ \text{J kg}^{-1}\ \text{K}^{-1}}$$