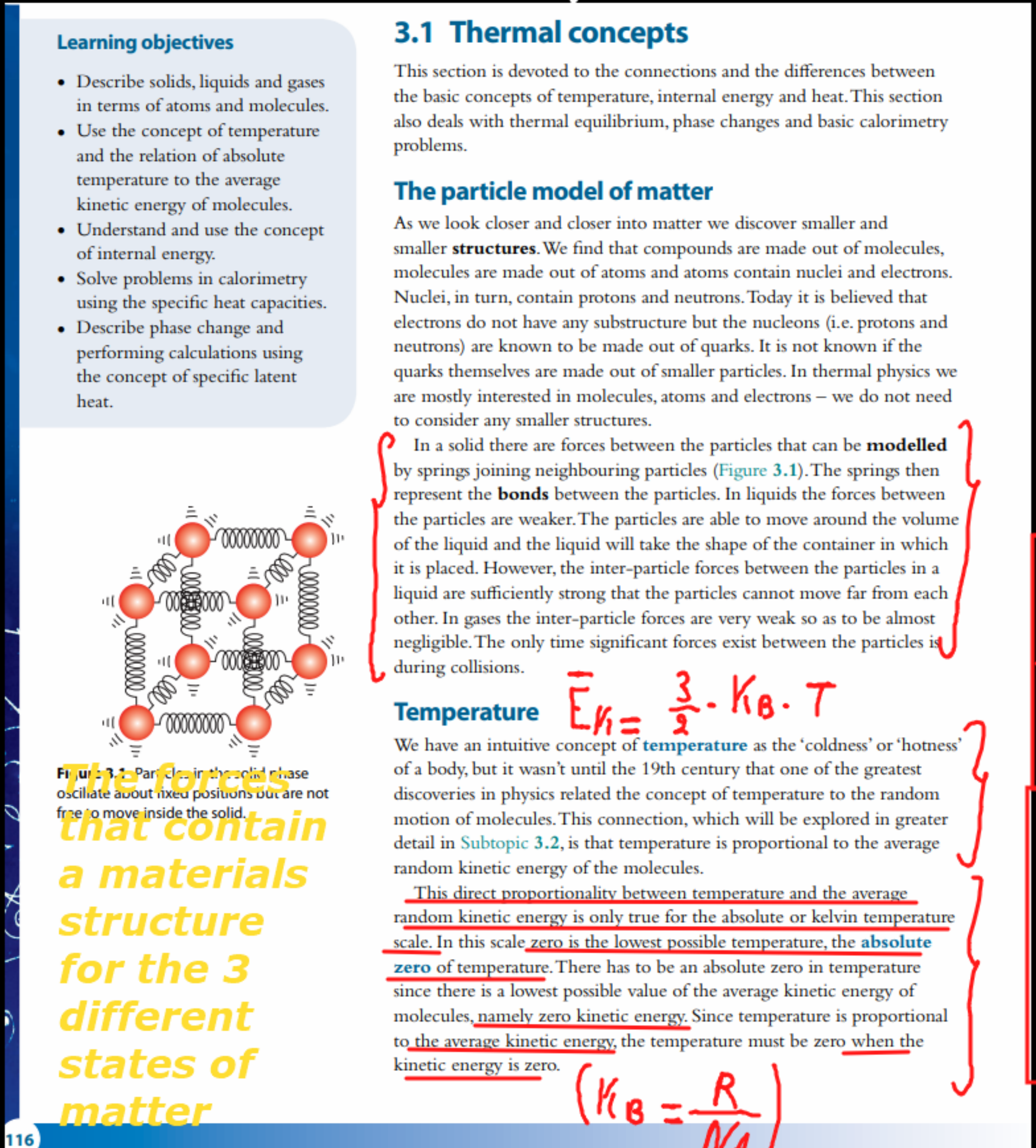
Solid (S)
Liquid (Li)
Gas (g)
states of matter
(Phase - Change of Phase)
State of matter you're currently on.



Learning objectives

- Describe solids, liquids and gases in terms of atoms and molecules.
- Use the concept of temperature and the relation of absolute temperature to the average kinetic energy of molecules.
- Understand and use the concept of internal energy.
- Solve problems in calorimetry using the specific heat capacities.
- Describe phase change and performing calculations using the concept of specific latent heat.

3.1 Thermal concepts

This section is devoted to the connections and the differences between the basic concepts of temperature, internal energy and heat. This section also deals with thermal equilibrium, phase changes and basic calorimetry problems.

The particle model of matter

As we look closer and closer into matter we discover smaller and smaller **structures**. We find that compounds are made out of molecules, molecules are made out of atoms and atoms contain nuclei and electrons. Nuclei, in turn, contain protons and neutrons. Today it is believed that electrons do not have any substructure but the nucleons (i.e. protons and neutrons) are known to be made out of quarks. It is not known if the quarks themselves are made out of smaller particles. In thermal physics we are mostly interested in molecules, atoms and electrons – we do not need to consider any smaller structures.

In a solid there are forces between the particles that can be **modelled** by springs joining neighbouring particles (Figure 3.1). The springs then represent the **bonds** between the particles. In liquids the forces between the particles are weaker. The particles are able to move around the volume of the liquid and the liquid will take the shape of the container in which it is placed. However, the inter-particle forces between the particles in a liquid are sufficiently strong that the particles cannot move far from each other. In gases the inter-particle forces are very weak so as to be almost negligible. The only time significant forces exist between the particles is during collisions.

Figure 3.1 Particles in the solid phase oscillate about fixed positions but are not free to move inside the solid.

The forces that contain a materials structure for the 3 different states of matter

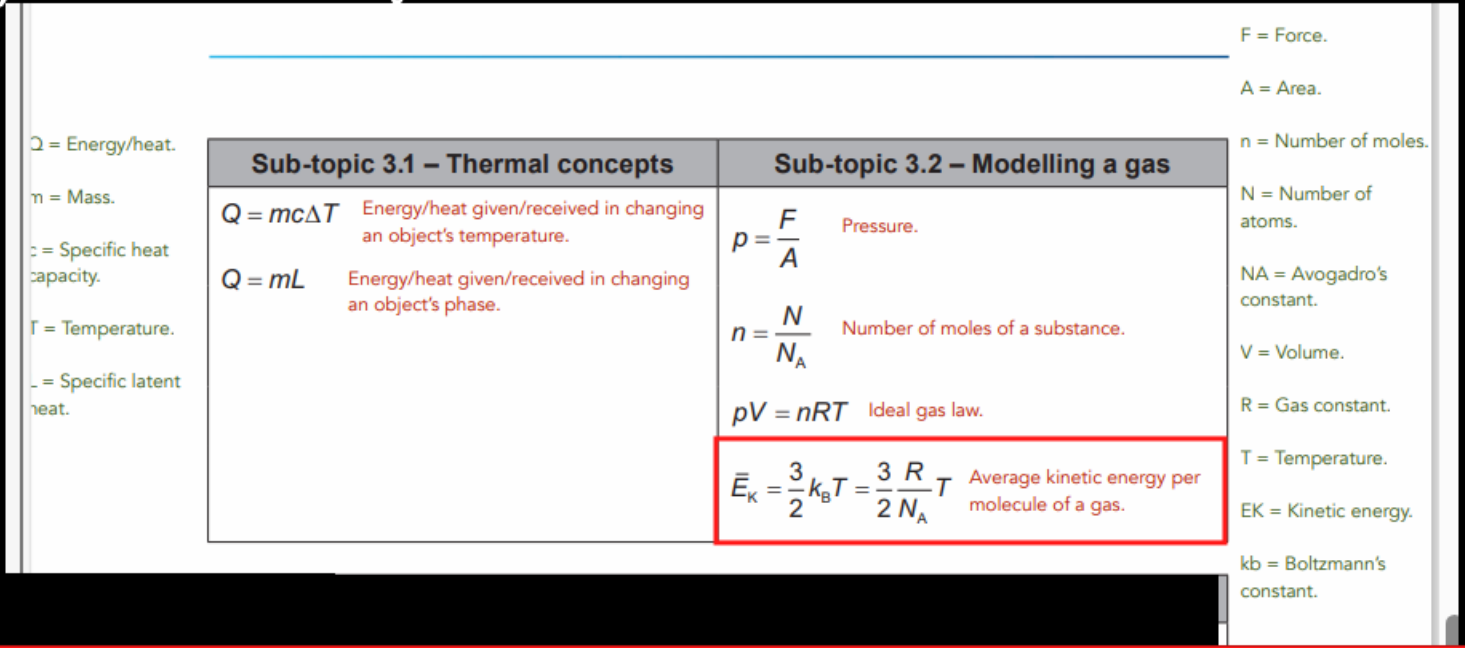
Temperature

$\bar{E}_K = \frac{3}{2} \cdot K_B \cdot T$

We have an intuitive concept of **temperature** as the 'coldness' or 'hotness' of a body, but it wasn't until the 19th century that one of the greatest discoveries in physics related the concept of temperature to the random motion of molecules. This connection, which will be explored in greater detail in Subtopic 3.2, is that temperature is proportional to the average random kinetic energy of the molecules.

This direct proportionality between temperature and the average random kinetic energy is only true for the absolute or kelvin temperature scale. In this scale zero is the lowest possible temperature, the **absolute zero** of temperature. There has to be an absolute zero in temperature since there is a lowest possible value of the average kinetic energy of molecules, namely zero kinetic energy. Since temperature is proportional to the average kinetic energy, the temperature must be zero when the kinetic energy is zero.

$(K_B = \frac{R}{N_A}$



Q = Energy/heat.
m = Mass.
c = Specific heat capacity.
T = Temperature.
L = Specific latent heat.

Sub-topic 3.1 – Thermal concepts		Sub-topic 3.2 – Modelling a gas	
$Q = mc\Delta T$	Energy/heat given/received in changing an object's temperature.	$p = \frac{F}{A}$	Pressure.
$Q = mL$	Energy/heat given/received in changing an object's phase.	$n = \frac{N}{N_A}$	Number of moles of a substance.
		$pV = nRT$	Ideal gas law.
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The average Kinetic Energy Per Molecule of Gas is Proportional to the temperature of a gas (or a liquid).

The higher the temperature, the higher the average kinetic energy per molecule of a gas (or a liquid).

Kb is the boltzman constant, specifically it is a way to statistically describe a gas.

Furthermore, while using this formula we only refer to molecules since we're dealing absolute zero temperature.

Many other temperature scales exist. In 1742, Anders Celsius (1701–1744) created the temperature scale that is still commonly used today and is known by his name. On the Celsius scale a value of zero degrees is assigned to the freezing point of water (Figure 3.2) and a value of 100 degrees is assigned to the boiling point of water. The connection between the Celsius and Kelvin scales is:

$$T \text{ (in kelvin, K)} = T \text{ (in degrees Celsius, °C)} + 273$$

The magnitude of a kelvin is the same as that of a degree Celsius.

The lowest possible temperature on the absolute scale is zero kelvin, 0 K. On the Celsius scale the lowest possible temperature is, therefore, −273 °C. (Notice that we never say degrees kelvin, just kelvin.)

Temperature has varied a lot in the life of the Universe: at the time of the Big Bang, some 13.8 billion years ago, the temperature of the universe was about 10^{32} K. The Universe has been expanding ever since and so the temperature has been dropping. In the emptiness of space, far from stars and galaxies, its value today is only 2.7 K.

Exam tip

The magnitude of a kelvin is the same as that of a degree Celsius.

The need to agree on internationally accepted units, among them those for temperature, is a good example of international collaboration to establish international systems of measurement.

Worked example

3.1 The temperature of a body increases from 320 K to 340 K. State the temperature increase in degrees Celsius.

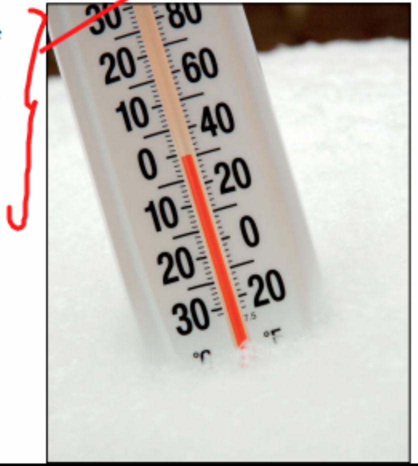
$(320 - 273)\text{K} = 47°\text{C}$
$(340 - 273)\text{K} = 67°\text{C}$

$\Delta T = 67 - 47 = 20°\text{C}$
final-initial
20°c increase.

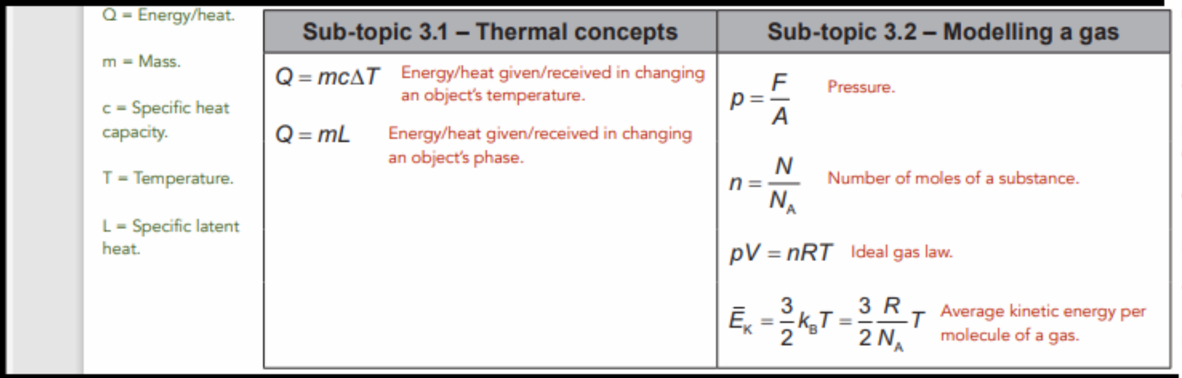
Measuring temperature

Temperature can be measured with a thermometer, which is simply a device that has one property that changes in a predictable way as temperature changes. That property is volume in liquid-in-glass thermometers: the liquid column changes its volume and hence its length since the cross-sectional area stays the same when the temperature changes and so can be used to measure temperature if we first **calibrate** the thermometer. But properties other than volume can be used, for example, electrical resistance.

When a thermometer is used to measure the temperature of a body it has to come into contact with the body. A **thermal interaction** takes place and energy is transferred until the thermometer and the body are at the same temperature. When this happens we say that we have **thermal equilibrium**. The reading on the thermometer is then the temperature of the body. (For thermometers such as infrared thermometers thermal contact is not necessary – the thermometer absorbs radiation emitted by the body whose temperature is to be measured.) The average temperature



Heat (Q) is a form of energy and it's not the same thing with Temperature (T).



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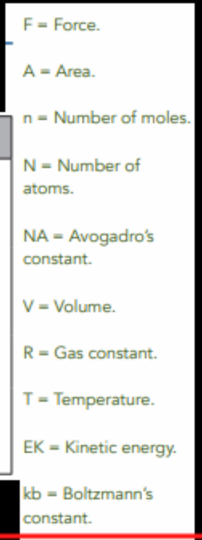

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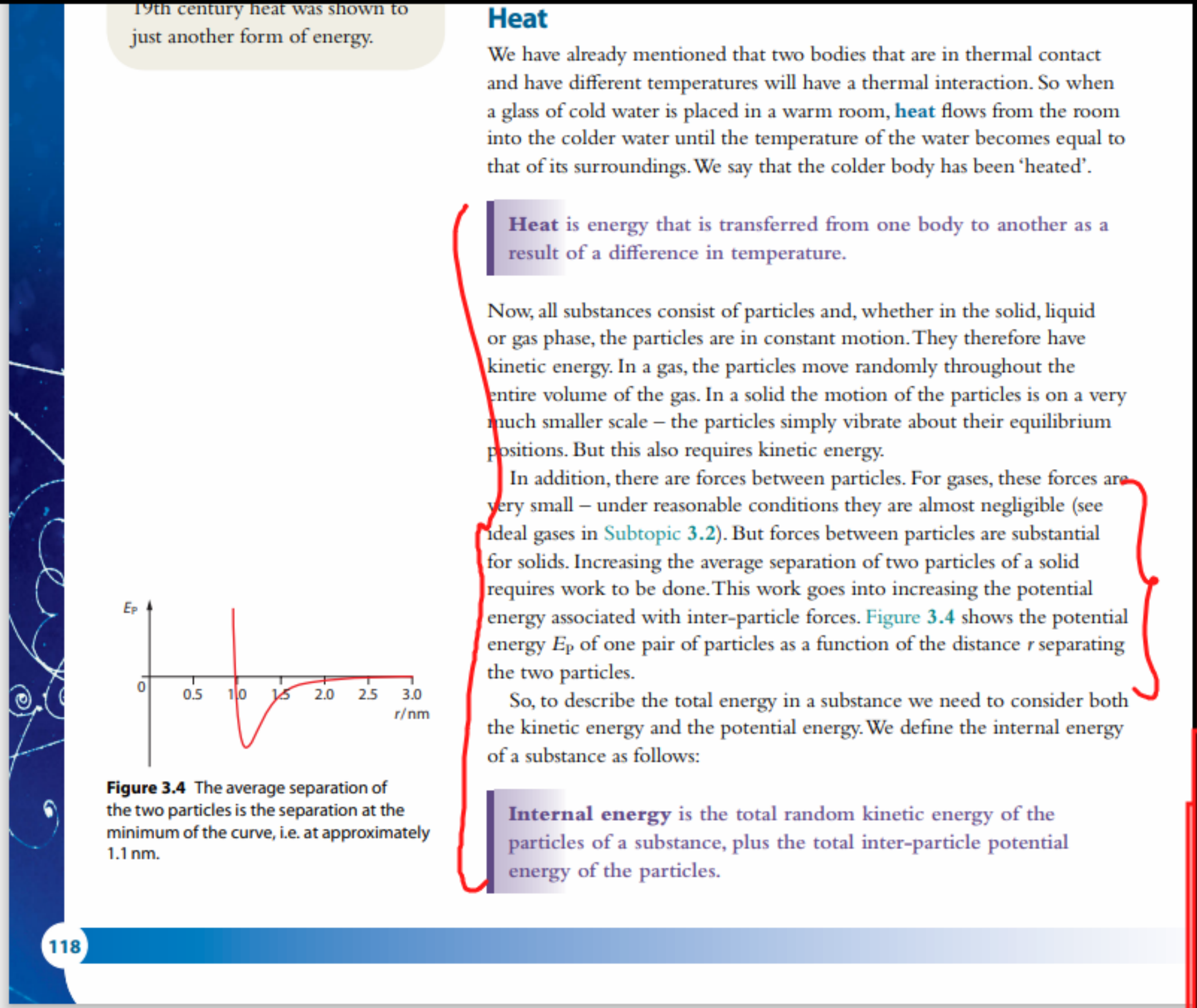
This is practically the science behind electrical sensors. The basic idea is that we relate a physical property (like volume of the liquid or electrical resistance) to the desired physical property that needs to be measured.

Thermal Equilibrium: 2 Different objects come into contact, they have different temperatures but after a while, since energy is transferred from the hotter object to the colder, we have the same temperature in both objects and thus "thermal equilibrium".
Exothermal Reaction: Heat is produced from the system and is transferred to the enviroment
Endothermal Reaction: Heat is provided to the system from the enviroment



19th century heat was shown to just another form of energy.

Heat

We have already mentioned that two bodies that are in thermal contact and have different temperatures will have a thermal interaction. So when a glass of cold water is placed in a warm room, **heat** flows from the room into the colder water until the temperature of the water becomes equal to that of its surroundings. We say that the colder body has been 'heated'.

> **Heat** is energy that is transferred from one body to another as a result of a difference in temperature.

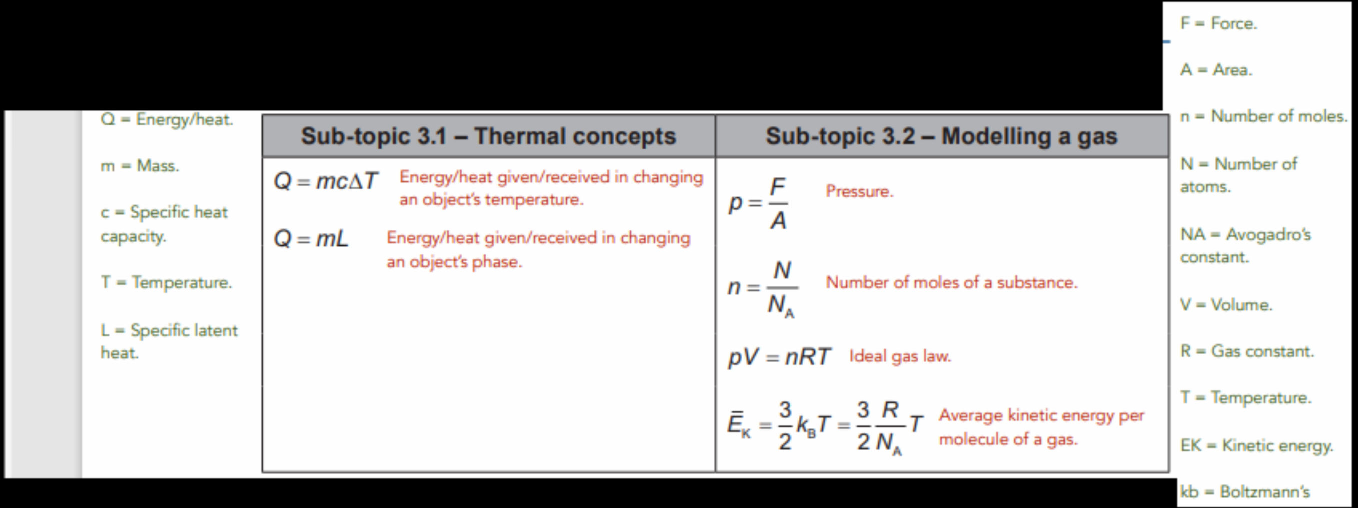
Now, all substances consist of particles and, whether in the solid, liquid or gas phase, the particles are in constant motion. They therefore have kinetic energy. In a gas, the particles move randomly throughout the entire volume of the gas. In a solid the motion of the particles is on a very much smaller scale – the particles simply vibrate about their equilibrium positions. But this also requires kinetic energy.

In addition, there are forces between particles. For gases, these forces are very small – under reasonable conditions they are almost negligible (see ideal gases in Subtopic 3.2). But forces between particles are substantial for solids. Increasing the average separation of two particles of a solid requires work to be done. This work goes into increasing the potential energy associated with inter-particle forces. Figure 3.4 shows the potential energy E_P of one pair of particles as a function of the distance r separating the two particles.

So, to describe the total energy in a substance we need to consider both the kinetic energy and the potential energy. We define the internal energy of a substance as follows:

> **Internal energy** is the total random kinetic energy of the particles of a substance, plus the total inter-particle potential energy of the particles.

Figure 3.4 The average separation of the two particles is the separation at the minimum of the curve, i.e. at approximately 1.1 nm.



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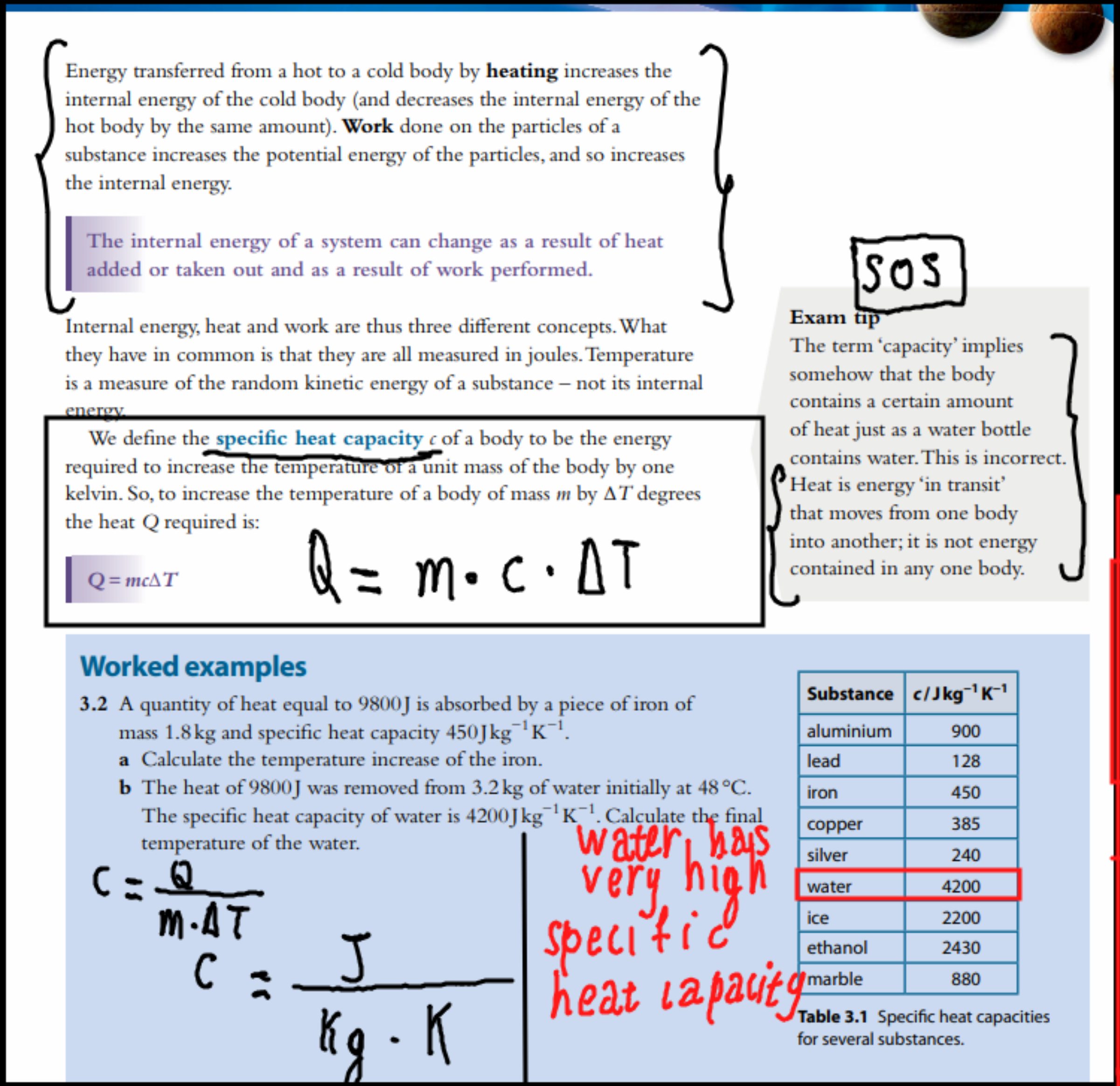
Particles have both kinetic and potential energy, and when you have those 2 combined you call this energy "Internal Energy". So it is kind of the "mechanical energy of the particles".

Particles have kinetic energy because they are in constant motion and they have potential energy due to the interparticle forces that need to exist in order for the particles of the materials to be bonded (to be connected or even better to remain one object)



Energy transferred from a hot to a cold body by **heating** increases the internal energy of the cold body (and decreases the internal energy of the hot body by the same amount). **Work** done on the particles of a substance increases the potential energy of the particles, and so increases the internal energy.

> The internal energy of a system can change as a result of heat added or taken out and as a result of work performed.

Internal energy, heat and work are thus three different concepts. What they have in common is that they are all measured in joules. Temperature is a measure of the random kinetic energy of a substance – not its internal energy.

We define the **specific heat capacity** c of a body to be the energy required to increase the temperature of a unit mass of the body by one kelvin. So, to increase the temperature of a body of mass m by ΔT degrees the heat Q required is:

$Q = mc\Delta T$

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

SOS

Exam tip

The term 'capacity' implies somehow that the body contains a certain amount of heat just as a water bottle contains water. This is incorrect. Heat is energy 'in transit' that moves from one body into another; it is not energy contained in any one body.

Worked examples

3.2 A quantity of heat equal to 9800 J is absorbed by a piece of iron of mass 1.8 kg and specific heat capacity 450 J kg^{-1}K^{-1}.

- a Calculate the temperature increase of the iron.
- b The heat of 9800 J was removed from 3.2 kg of water initially at 48 °C. The specific heat capacity of water is 4200 J kg^{-1}K^{-1}. Calculate the final temperature of the water.

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

$$c = \frac{J}{Kg \cdot K}$$

water has very high specific heat capacity

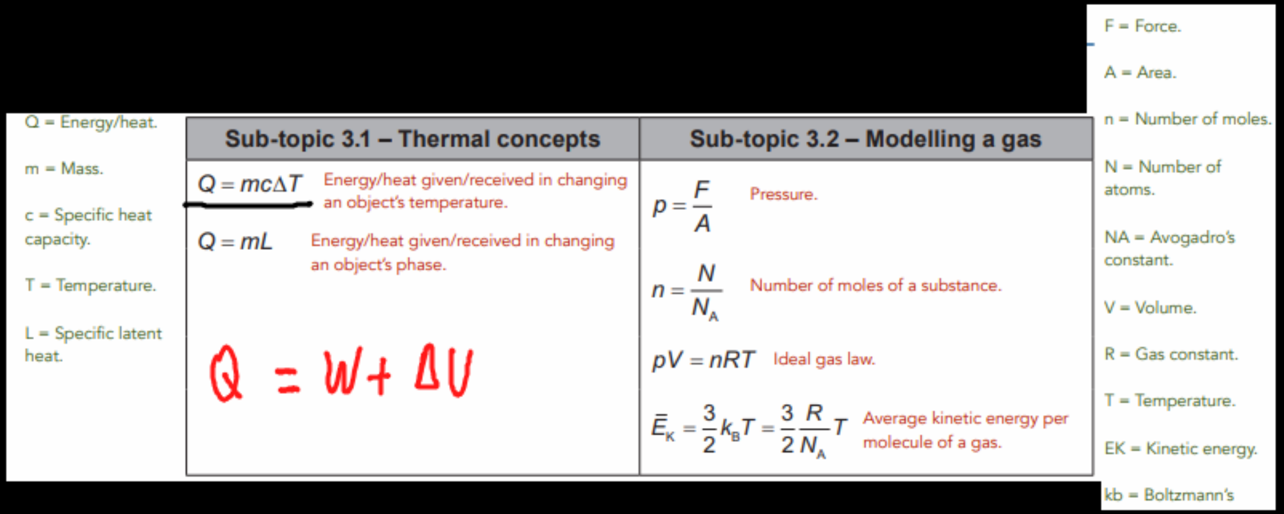
Substance	c / J kg^{-1} K^{-1}
aluminium	900
lead	128
iron	450
copper	385
silver	240
water	4200
ice	2200
ethanol	2430
marble	880

Table 3.1 Specific heat capacities for several substances.

ⓐ
$Q = 9{,}800\,J$
$m = 1.8\,Kg$
$c = 450\,J \cdot Kg^{-1} \cdot K^{-1}$

$$\Delta T = \frac{Q}{m \cdot c} \Leftrightarrow \boxed{\Delta T = 12.1K}$$

$$\Delta T = \frac{9{,}800\,\cancel{J}}{1.8\,\cancel{Kg} \cdot 450 \cdot \cancel{J} \cdot \cancel{Kg}^{-1} \cdot K^{-1}} \Rightarrow$$



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First Law of Thermodynamics: Q = W + ΔU

The heat that is produced in an event (reaction) and that is provided or extracted from a thermodynamic system is equal to the sum of the work done (from the system) + the change in the internal energy of the system

Special Cases (Ideal Gases)

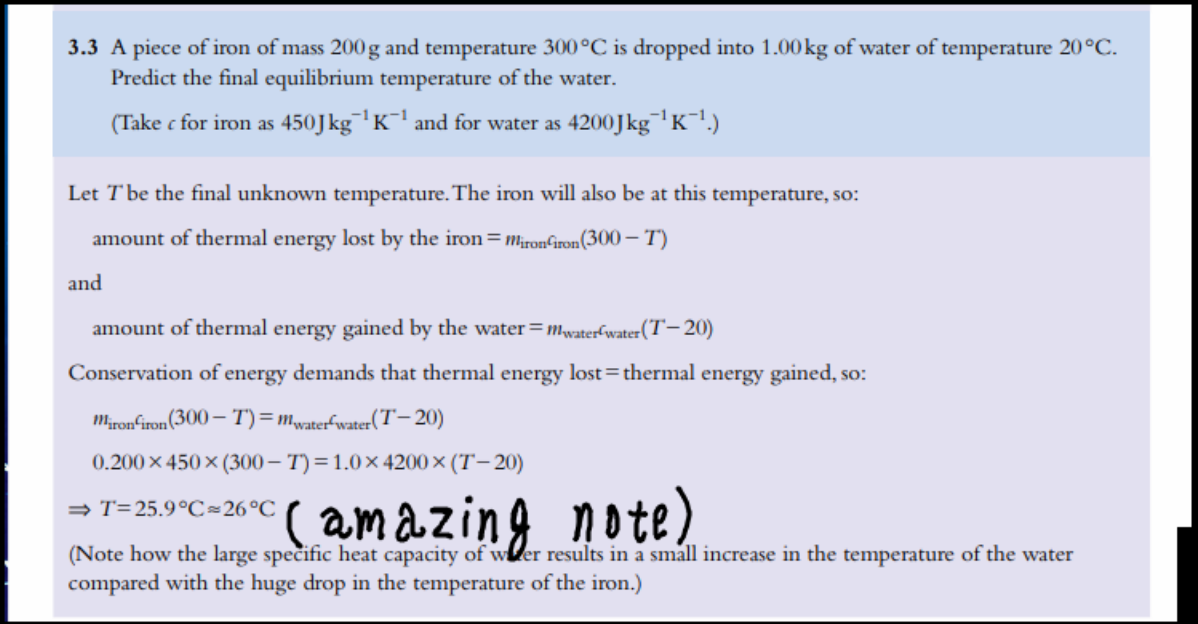
$\Delta U = 0$ because $(\Delta T = 0)$

① When you have no change in temperature: $\Delta U = \frac{3}{2} n \cdot R \cdot \underset{0}{\Delta T} = 0$
$Q = W \Rightarrow$ The whole heat transforms into "beneficial work"

② When $Q = 0J \Rightarrow W = -\Delta U$ (no heat transfer or extraction to the enviroment)

③ When $W = 0J \Rightarrow Q = \Delta U$ (no work produced)

ⓑ $\Delta T = \frac{Q}{m \cdot c} \Leftrightarrow T_f - T_{ini} = \frac{Q}{m \cdot c} \Leftrightarrow T_f = \frac{Q}{m \cdot c} + T_{ini} =$ K.



3.3 A piece of iron of mass 200 g and temperature 300 °C is dropped into 1.00 kg of water of temperature 20 °C. Predict the final equilibrium temperature of the water.

(Take c for iron as 450 J kg^{-1} K^{-1} and for water as 4200 J kg^{-1} K^{-1}.)

Let T be the final unknown temperature. The iron will also be at this temperature, so:

amount of thermal energy lost by the iron $= m_{\text{iron}}c_{\text{iron}}(300 - T)$

and

amount of thermal energy gained by the water $= m_{\text{water}}c_{\text{water}}(T - 20)$

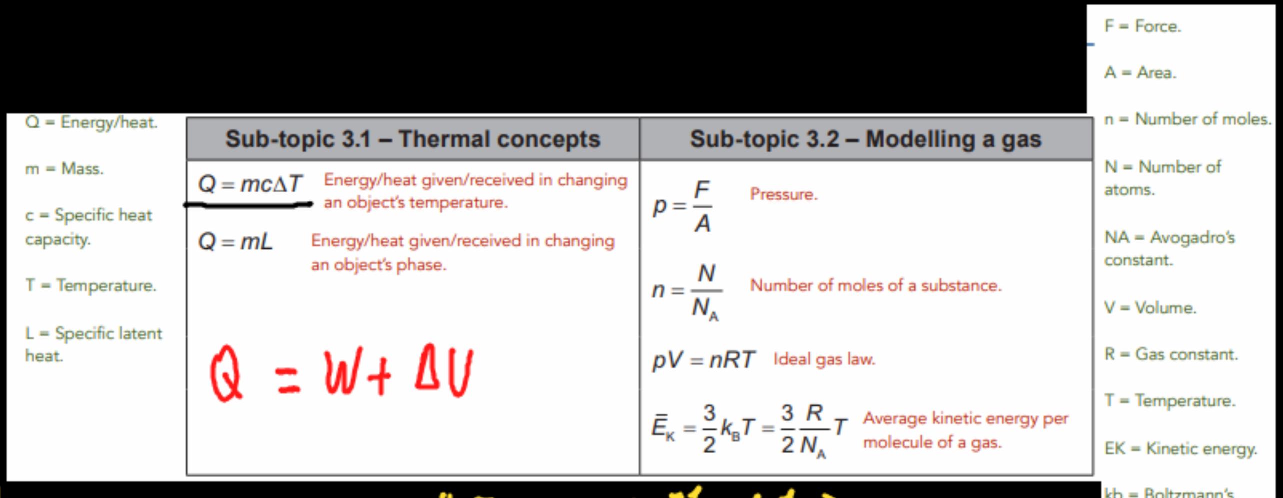
Conservation of energy demands that thermal energy lost = thermal energy gained, so:

$$m_{\text{iron}}c_{\text{iron}}(300 - T) = m_{\text{water}}c_{\text{water}}(T - 20)$$

$$0.200 \times 450 \times (300 - T) = 1.0 \times 4200 \times (T - 20)$$

$\Rightarrow T = 25.9\,°\text{C} \approx 26\,°\text{C}$ (amazing note)

(Note how the large specific heat capacity of water results in a small increase in the temperature of the water compared with the huge drop in the temperature of the iron.)



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$$\left.\begin{array}{l} m_i = 0.2\ \text{Kg} \mid m_{w.} = 1\ \text{Kg} \\ T_{ir.} = 300°C = 573\ K \\ T_{w.} = 293\ K \end{array}\right\} \quad \left.\begin{array}{l} C_i = 450\ J \cdot Kg^{-1} \cdot K^{-1} \\ C_w = 4{,}200\ J\ Kg^{-1} \cdot K^{-1} \end{array}\right\}$$

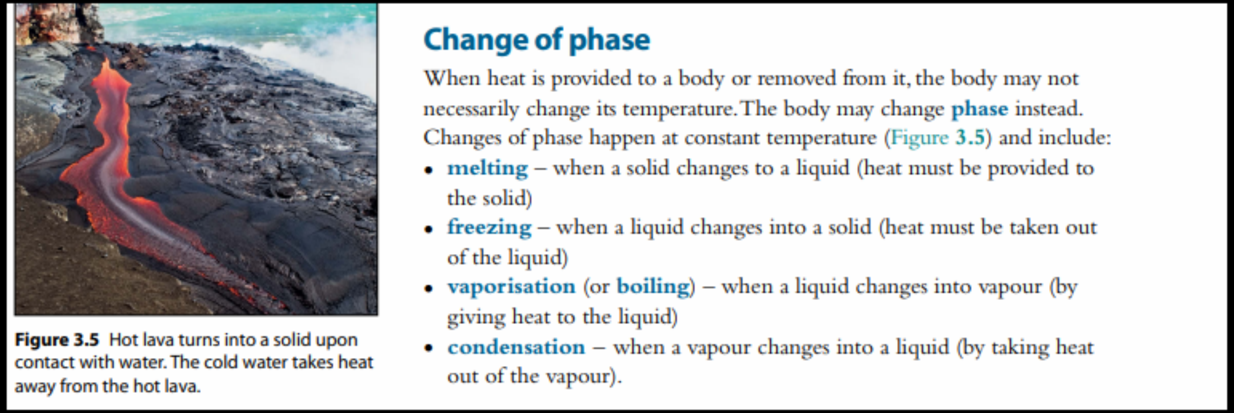
3.3

In newtonian mechanics, we use the term "conservation of mechanical energy", in thermodynamics, energy is also conserved, thus we can practically refer to it as "concervation of heat - conservation of thermal energy".

We have to express the 2 heats, specifically we have an "iron - water" system. The iron gives heat to system while the water receives the heat.

- Let T_f, be the final of their equilibrium. $Q_w = Q_{ir} \Leftrightarrow m_w \cdot C_w \cdot (T_f - 293) = m_{ir.} \cdot C_{ir.} \cdot (T_f - 573) \Leftrightarrow$

$1 \cdot 4{,}200 \cdot (T_f - 293) = 0.2 \cdot 450 \cdot (T_f - 573) \Leftrightarrow \boxed{T_f = 299\ K} \sim 26°C$

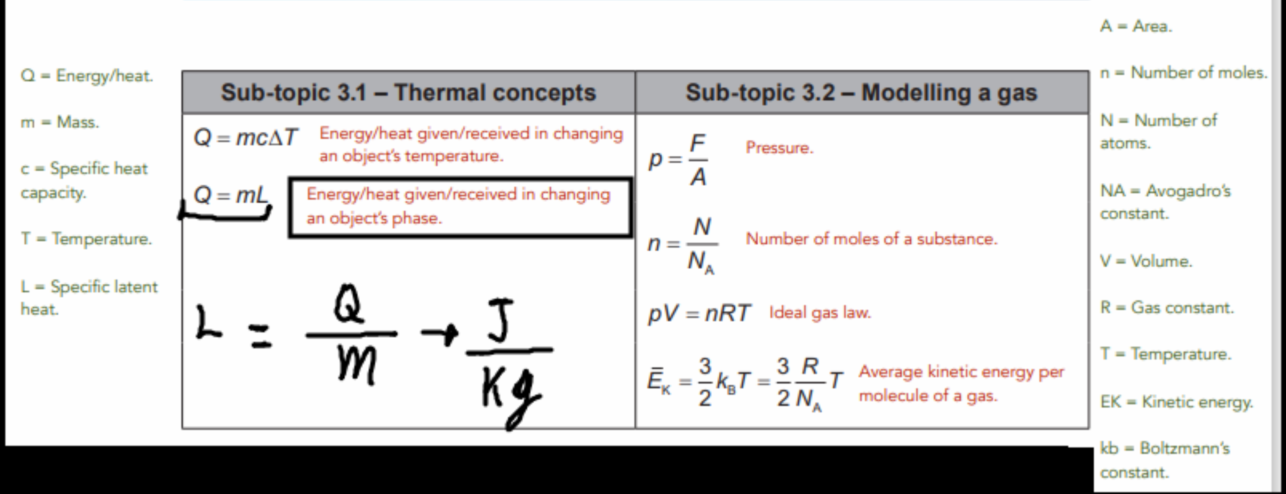


Change of phase

When heat is provided to a body or removed from it, the body may not necessarily change its temperature. The body may change **phase** instead. Changes of phase happen at constant temperature (Figure 3.5) and include:

- **melting** – when a solid changes to a liquid (heat must be provided to the solid)
- **freezing** – when a liquid changes into a solid (heat must be taken out of the liquid)
- **vaporisation** (or **boiling**) – when a liquid changes into vapour (by giving heat to the liquid)
- **condensation** – when a vapour changes into a liquid (by taking heat out of the vapour).

Figure 3.5 Hot lava turns into a solid upon contact with water. The cold water takes heat away from the hot lava.



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$\underline{Q = mL}$	$\boxed{\text{Energy/heat given/received in changing an object's phase.}}$	$n = \frac{N}{N_A}$	Number of moles of a substance.
$L = \frac{Q}{m} \rightarrow \frac{J}{Kg}$		$pV = nRT$	Ideal gas law.
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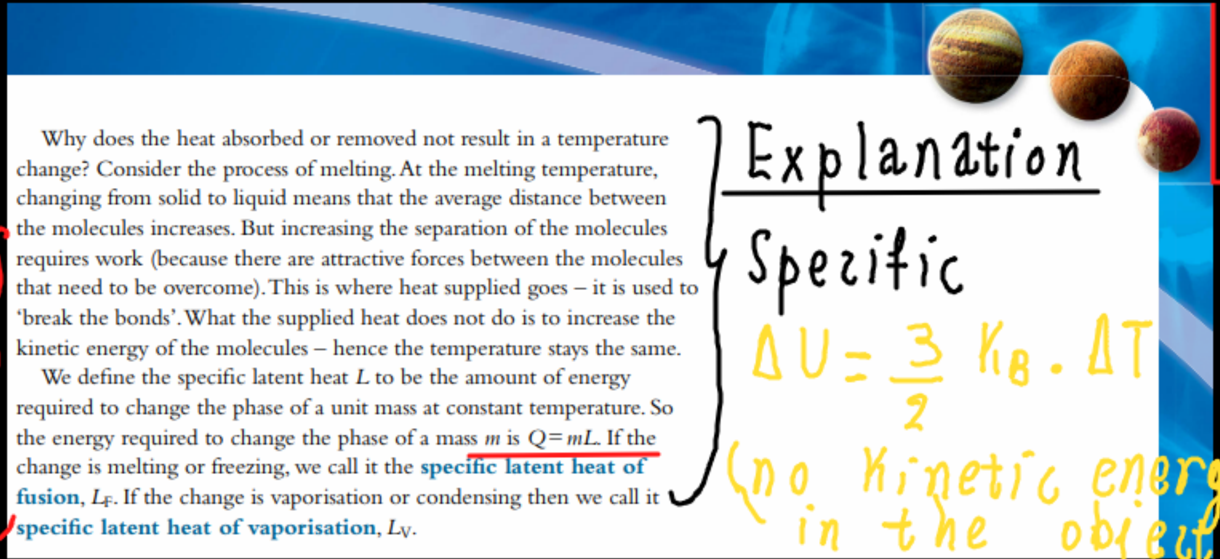


Why does the heat absorbed or removed not result in a temperature change? Consider the process of melting. At the melting temperature, changing from solid to liquid means that the average distance between the molecules increases. But increasing the separation of the molecules requires work (because there are attractive forces between the molecules that need to be overcome). This is where heat supplied goes – it is used to 'break the bonds'. What the supplied heat does not do is to increase the kinetic energy of the molecules – hence the temperature stays the same.

We define the specific latent heat *L* to be the amount of energy required to change the phase of a unit mass at constant temperature. So the energy required to change the phase of a mass *m* is $Q = mL$. If the change is melting or freezing, we call it the **specific latent heat of fusion**, L_F. If the change is vaporisation or condensing then we call it **specific latent heat of vaporisation**, L_V.

Explanation

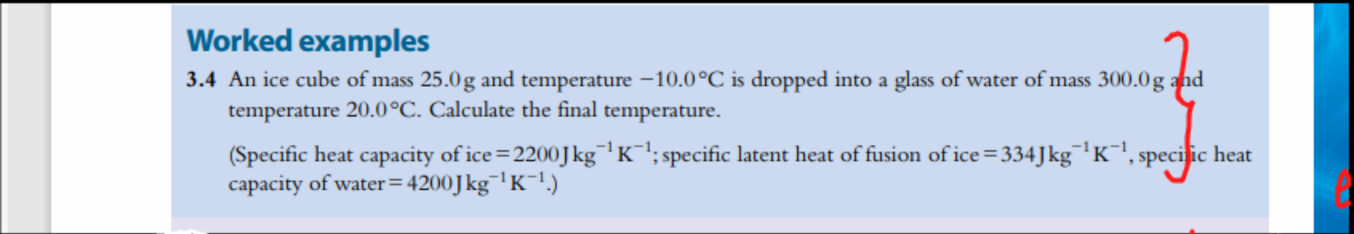
Spezific

$\Delta U = \frac{3}{2} K_B \cdot \Delta T$

(no kinetic energy provided to the sytem means that we have no change) in the object temperature.

Extremely Important Note: When we have a change in phase, temperature remains the same. Changes of phase happen at constant temperature.

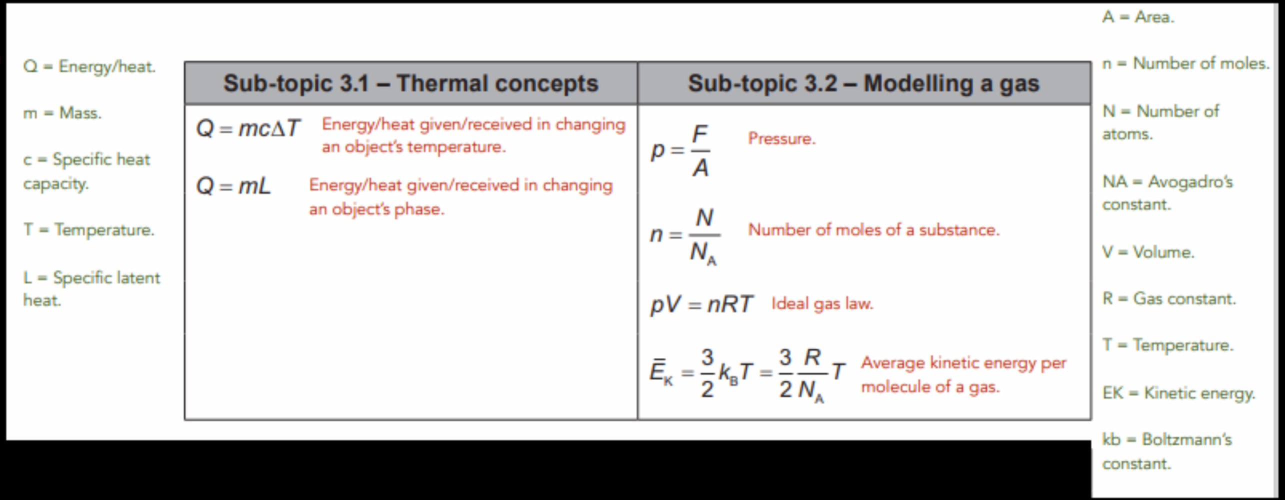
All the heat provided to the system (or absorbed from it) in order to have a change of phase (the state of matter of the object) is going to be used in order to "break the bonds" in order to overcome the interparticle forces used to hold the molecules of the substance together. So all the heat is going to be used to counter-act the potential energy stored within the material.



Worked examples

3.4 An ice cube of mass 25.0 g and temperature −10.0 °C is dropped into a glass of water of mass 300.0 g and temperature 20.0 °C. Calculate the final temperature.

(Specific heat capacity of ice = 2200 J kg^{-1} K^{-1}; specific latent heat of fusion of ice = 334 J kg^{-1} K^{-1}, specific heat capacity of water = 4200 J kg^{-1} K^{-1}.)



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We have to ignore the energy lost by the glass.

Our system consists of the ice cube and the water.

$\boxed{3.4}$ $\begin{cases} m_c = 0.025\ Kg \\ T_c = 263\ K \end{cases}$ $\begin{cases} m_w = 0.3\ Kg \\ T_w = 293 K \end{cases}$

Let T_f, be the final temperature

$Q_w = m_w \cdot C_w \cdot \Delta T$ (thermal energy lost by the water)

$Q_w = 0.3 \cdot 4{,}200 \cdot (T_f - 293)$ → to the ice → Q required for the change of phase

heat

$1{,}260\,(T_f - 293) = Q_{ice} + Q_{phase} + Q_{fin}$ → heat required for the system to reach its final temperature (Homework).

↑ required to raise the temperature of ice from 263K to 273K.

✓ **3.5** A sample of 120 g of a solid initially at 20 °C is heated by a heater of constant power. The specific heat capacity of the solid is 2500 J kg^{-1} K^{-1}. The temperature of the sample varies with time as shown in Figure 3.6.

Use the graph to determine:

- a the power of the heater
- b the melting temperature of the sample
- c the specific latent heat of fusion of the sample
- d the specific heat capacity of the sample in the liquid phase.

ⓐ $P = \frac{W}{t} \Rightarrow P = \frac{Q}{t}$ (Watt)

$t = 120s.$

$\boxed{m = 0.12\,Kg}$

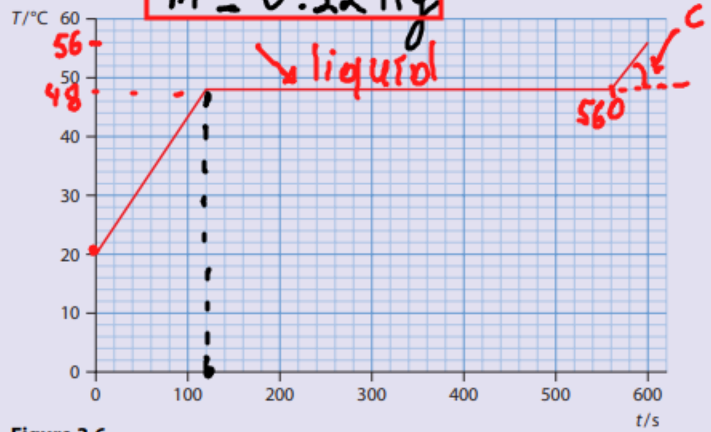


Figure 3.6

$T_1 = 20°C = 293K$
$T_2 = 48°C = 321K$

$\Delta T = 28K$

$Q = m \cdot c_s \cdot \Delta T = 0.12 \cdot 2{,}500 \cdot 28$

ⓐ $Q = 8400J \Rightarrow P = \frac{Q}{t} = \frac{840\cancel{0}}{12\cancel{0}} \Rightarrow P = 70$

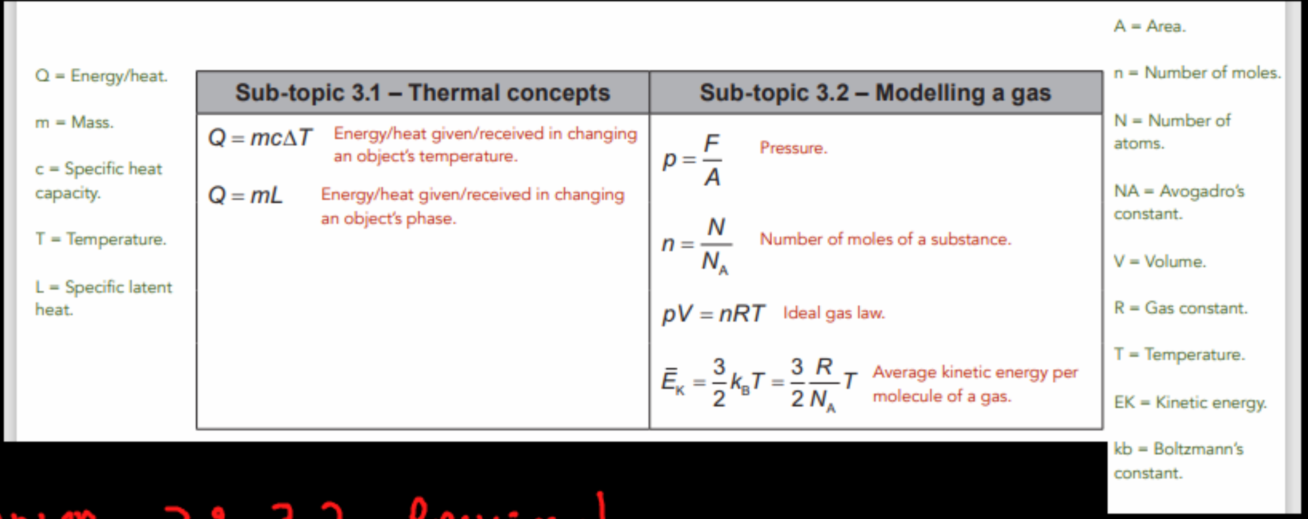
ⓑ $T_m = 48°C = 321K$

ⓒ $Q = m \cdot L \Rightarrow P \cdot t = m \cdot L \Leftrightarrow$

$t_f = (560 - 120)s \Leftrightarrow t = 440s$

$L_f = \frac{70 \cdot 440}{0.12} \Leftrightarrow L_f = 257.000\,J/Kg = 257\,KJ/Kg$

ⓓ $\Delta T = 8K$ $t = 40s$ $Q = m \cdot c \cdot \Delta T \Leftrightarrow c_L = \frac{P \cdot t}{m \cdot \Delta T} \Rightarrow c_L = 2{,}916\,J/Kg \cdot K$



Q = Energy/heat.
m = Mass.
c = Specific heat capacity.
T = Temperature.
L = Specific latent heat.

Sub-topic 3.1 – Thermal concepts		Sub-topic 3.2 – Modelling a gas	
$Q = mc\Delta T$	Energy/heat given/received in changing an object's temperature.	$p = \frac{F}{A}$	Pressure.
$Q = mL$	Energy/heat given/received in changing an object's phase.	$n = \frac{N}{N_A}$	Number of moles of a substance.
		$pV = nRT$	Ideal gas law.
		$\bar{E}_K = \frac{3}{2}k_B T = \frac{3}{2}\frac{R}{N_A}T$	Average kinetic energy per molecule of a gas.

A = Area.
n = Number of moles.
N = Number of atoms.
NA = Avogadro's constant.
V = Volume.
R = Gas constant.
T = Temperature.
EK = Kinetic energy.
kb = Boltzmann's constant.

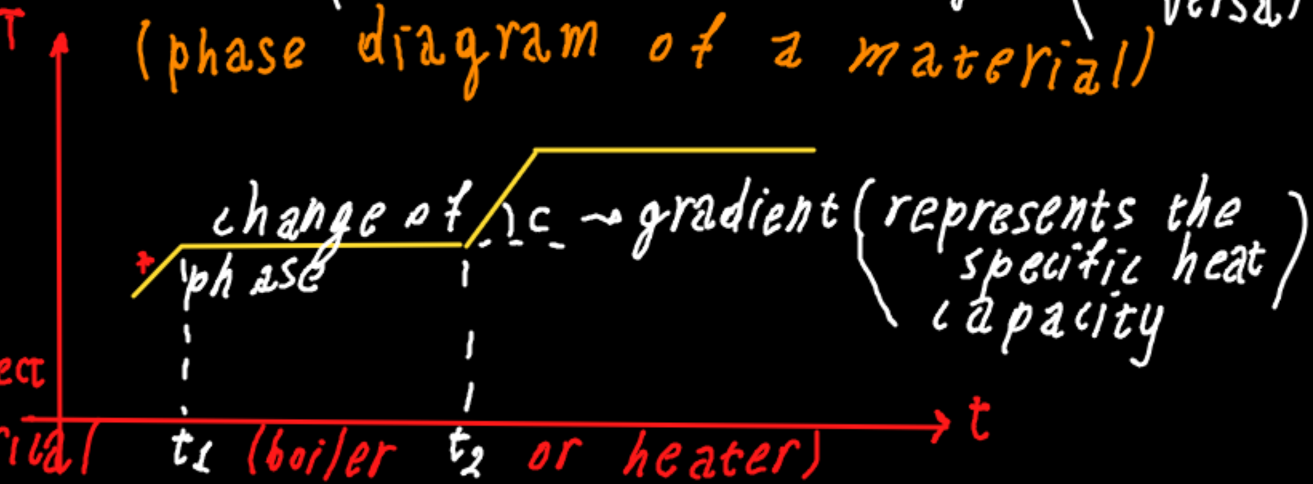
(Chapter 3.2-3.3 Preview)

Latent Heat: The energy required per kg (J/Kg) for the material under examination to change its phase.

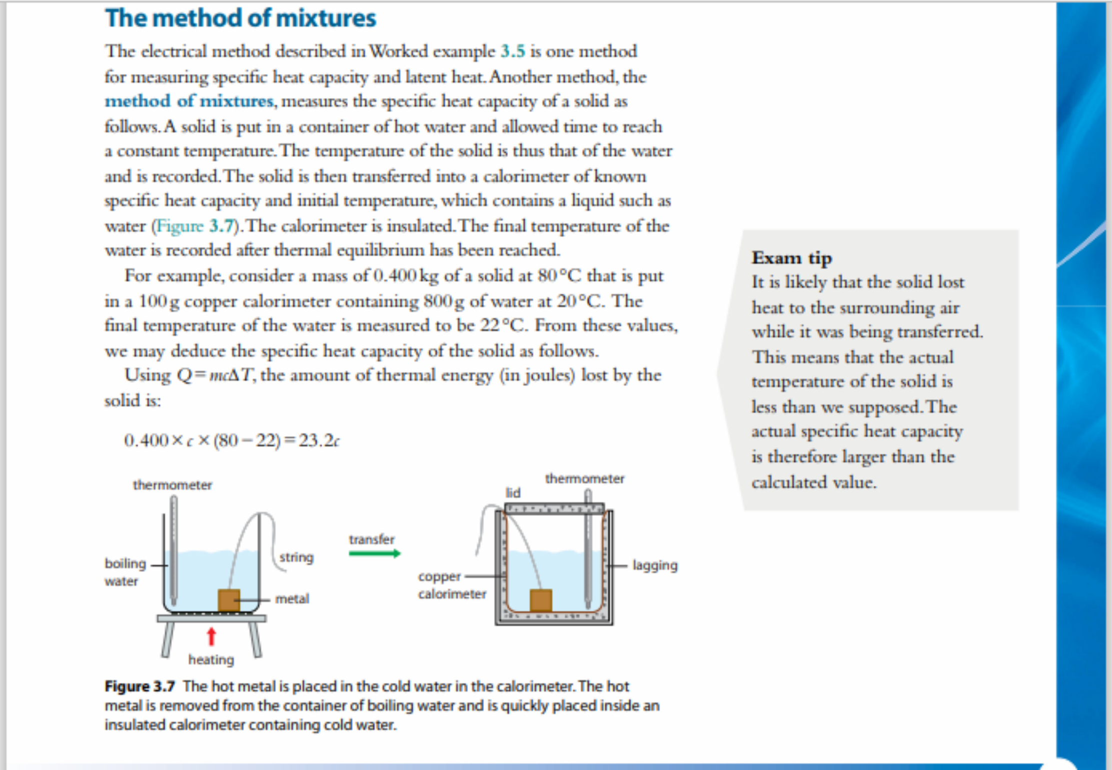
Latent Heat of fusion (L$_f$): Solid ⇌ Liquid (or vice versa)

Latent Heat of vaporization (L$_v$): Liquid ⇌ gas (or vice versa)

(phase diagram of a material)



under the effect of an electrical (boiler or heater)



The method of mixtures

The electrical method described in Worked example 3.5 is one method for measuring specific heat capacity and latent heat. Another method, the **method of mixtures**, measures the specific heat capacity of a solid as follows. A solid is put in a container of hot water and allowed time to reach a constant temperature. The temperature of the solid is thus that of the water and is recorded. The solid is then transferred into a calorimeter of known specific heat capacity and initial temperature, which contains a liquid such as water (Figure 3.7). The calorimeter is insulated. The final temperature of the water is recorded after thermal equilibrium has been reached.

For example, consider a mass of 0.400 kg of a solid at 80 °C that is put in a 100 g copper calorimeter containing 800 g of water at 20 °C. The final temperature of the water is measured to be 22 °C. From these values, we may deduce the specific heat capacity of the solid as follows.

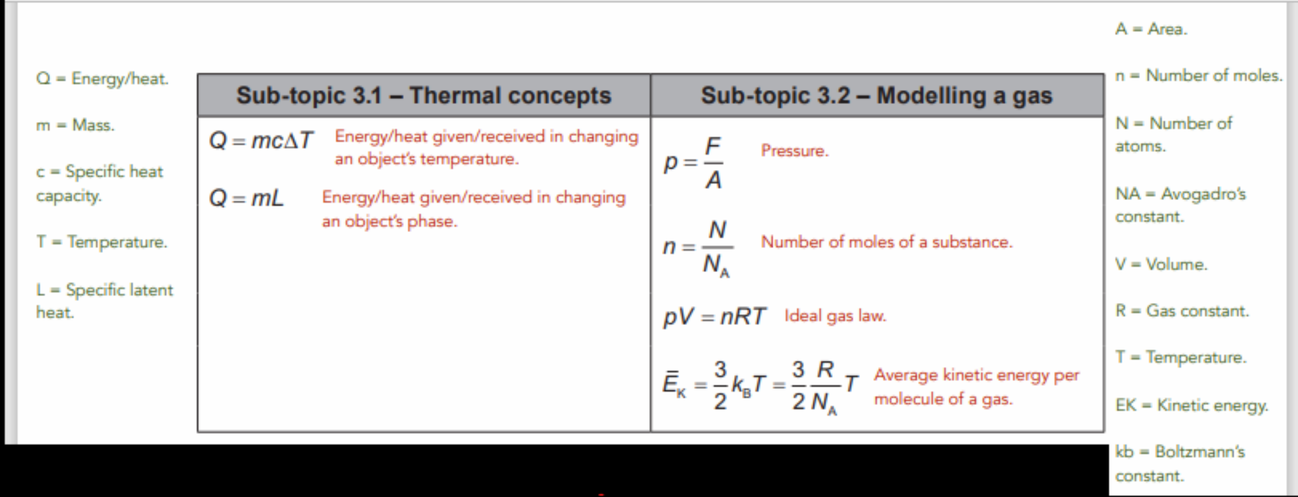
Using $Q = mc\Delta T$, the amount of thermal energy (in joules) lost by the solid is:

$$0.400 \times c \times (80 - 22) = 23.2c$$

Exam tip

It is likely that the solid lost heat to the surrounding air while it was being transferred. This means that the actual temperature of the solid is less than we supposed. The actual specific heat capacity is therefore larger than the calculated value.

Figure 3.7 The hot metal is placed in the cold water in the calorimeter. The hot metal is removed from the container of boiling water and is quickly placed inside an insulated calorimeter containing cold water.



Q = Energy/heat.

m = Mass.

c = Specific heat capacity.

T = Temperature.

L = Specific latent heat.

Sub-topic 3.1 – Thermal concepts		Sub-topic 3.2 – Modelling a gas	
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$Q = mL$	Energy/heat given/received in changing an object's phase.	$n = \frac{N}{N_A}$	Number of moles of a substance.
		$pV = nRT$	Ideal gas law.
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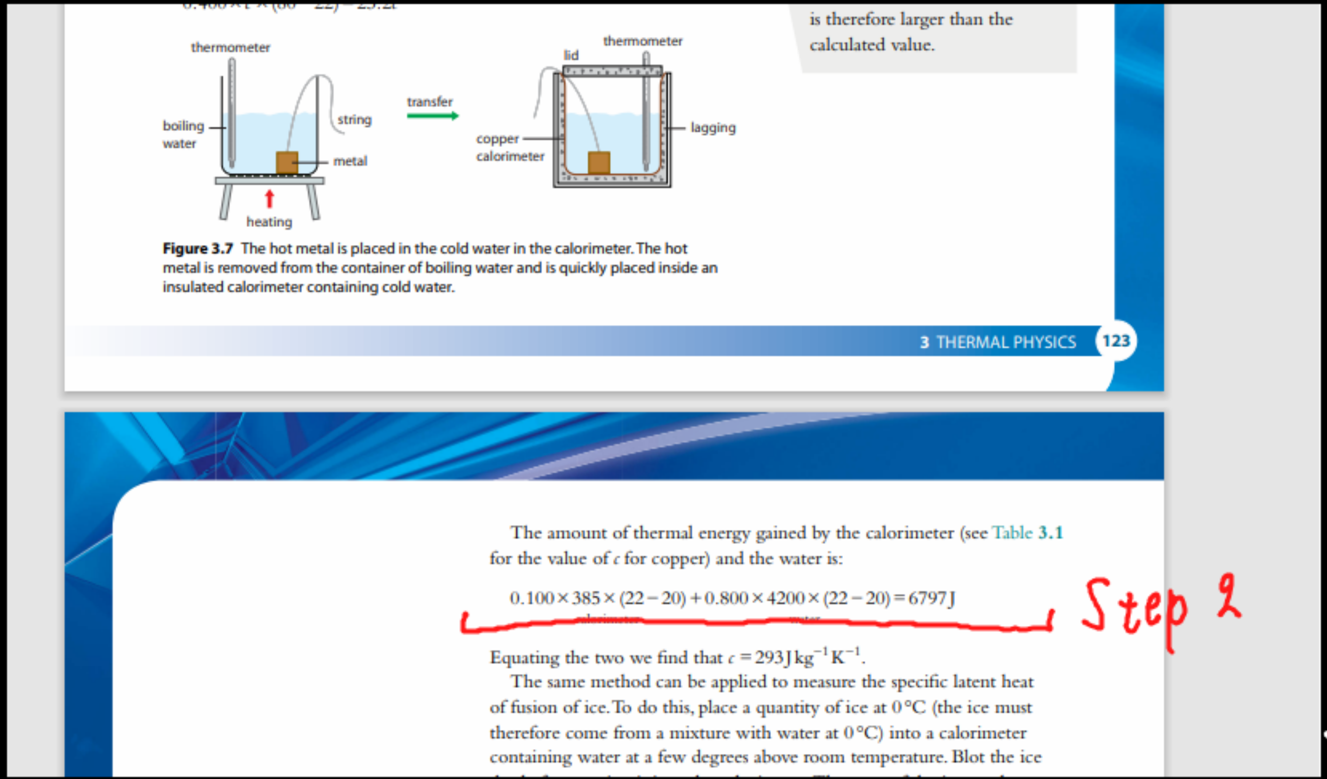


is therefore larger than the calculated value.

Figure 3.7 The hot metal is placed in the cold water in the calorimeter. The hot metal is removed from the container of boiling water and is quickly placed inside an insulated calorimeter containing cold water.

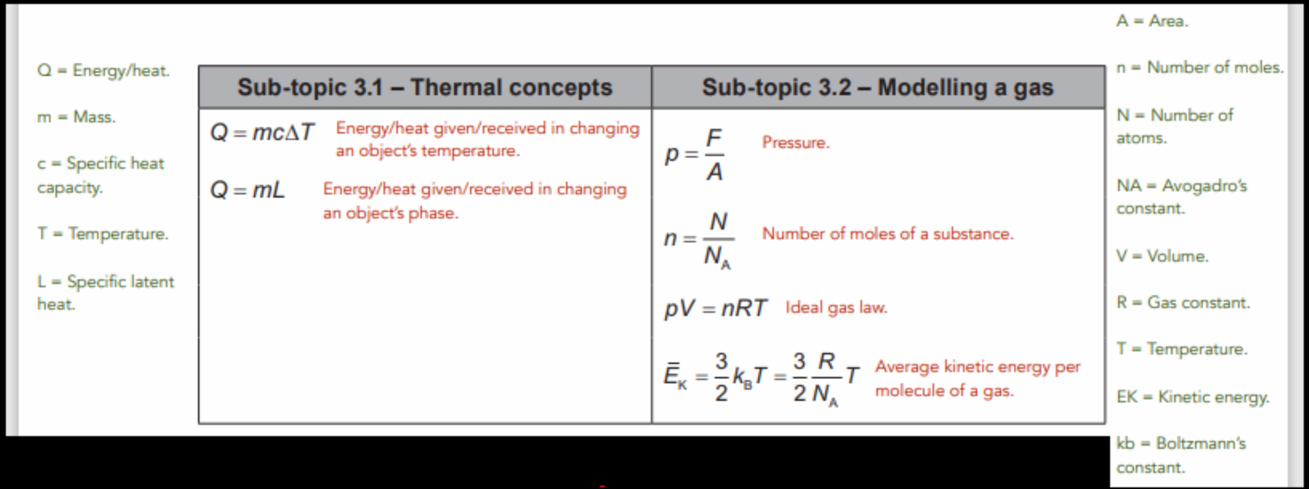
The amount of thermal energy gained by the calorimeter (see Table 3.1 for the value of c for copper) and the water is:

$$\underbrace{0.100\times385\times(22-20)}_{\text{calorimeter}}+\underbrace{0.800\times4200\times(22-20)}_{\text{water}}=6797\,\text{J}$$

Step 2

Equating the two we find that $c = 293\,\text{J kg}^{-1}\,\text{K}^{-1}$.

The same method can be applied to measure the specific latent heat of fusion of ice. To do this, place a quantity of ice at 0 °C (the ice must therefore come from a mixture with water at 0 °C) into a calorimeter containing water at a few degrees above room temperature. Blot the ice



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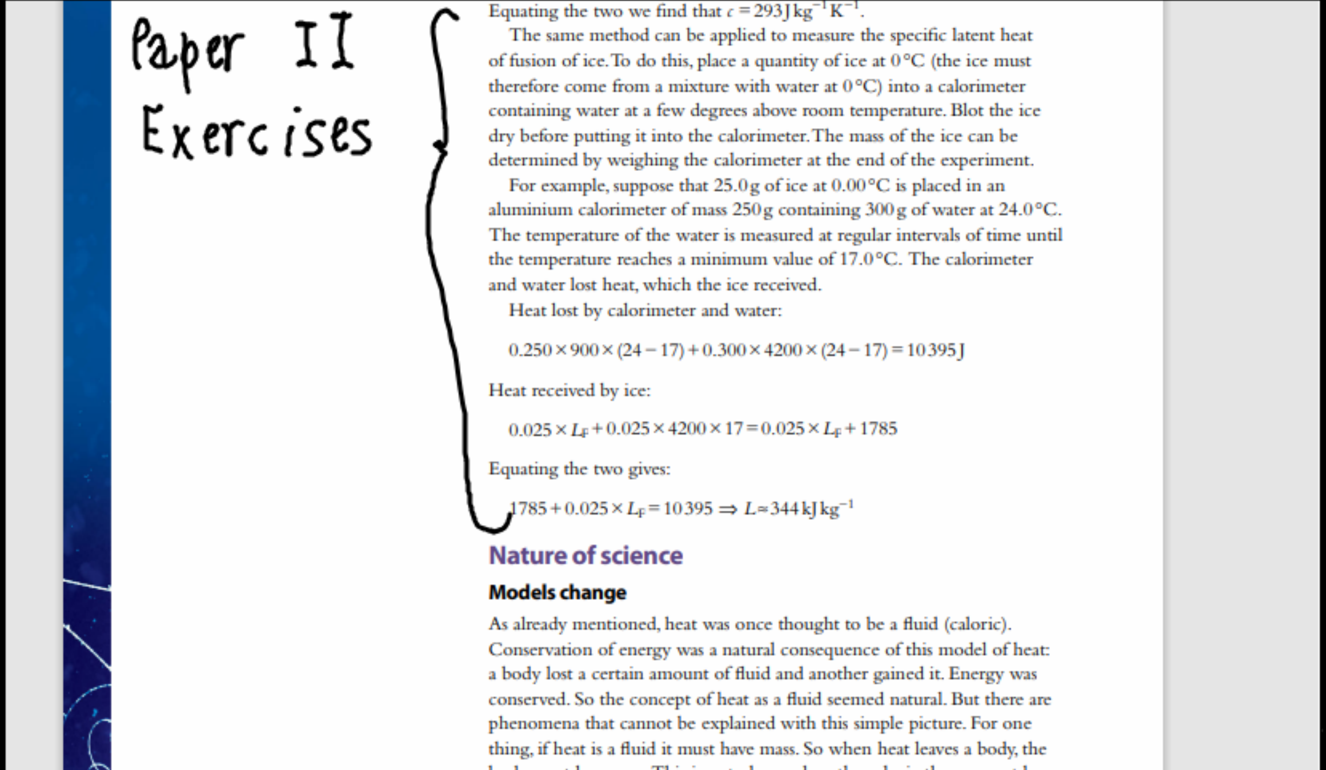
The Method of Mixtures:

Step 1: Create a starting equation with specific heat capacity as the unknown variable (from the first mixture)

Step 2: Calculate the total heat (Q) produced, without taking into account the object that we're trying to measure its specific heat capacity (c)

Step 3: Make the total heat measured equal to the equation that we have "produced" for the metallic object



Paper II
Exercises

Equating the two we find that $c = 293\,\mathrm{J\,kg^{-1}\,K^{-1}}$.

The same method can be applied to measure the specific latent heat of fusion of ice. To do this, place a quantity of ice at 0 °C (the ice must therefore come from a mixture with water at 0 °C) into a calorimeter containing water at a few degrees above room temperature. Blot the ice dry before putting it into the calorimeter. The mass of the ice can be determined by weighing the calorimeter at the end of the experiment.

For example, suppose that 25.0 g of ice at 0.00 °C is placed in an aluminium calorimeter of mass 250 g containing 300 g of water at 24.0 °C. The temperature of the water is measured at regular intervals of time until the temperature reaches a minimum value of 17.0 °C. The calorimeter and water lost heat, which the ice received.

Heat lost by calorimeter and water:

$$0.250 \times 900 \times (24 - 17) + 0.300 \times 4200 \times (24 - 17) = 10395\,\mathrm{J}$$

Heat received by ice:

$$0.025 \times L_F + 0.025 \times 4200 \times 17 = 0.025 \times L_F + 1785$$

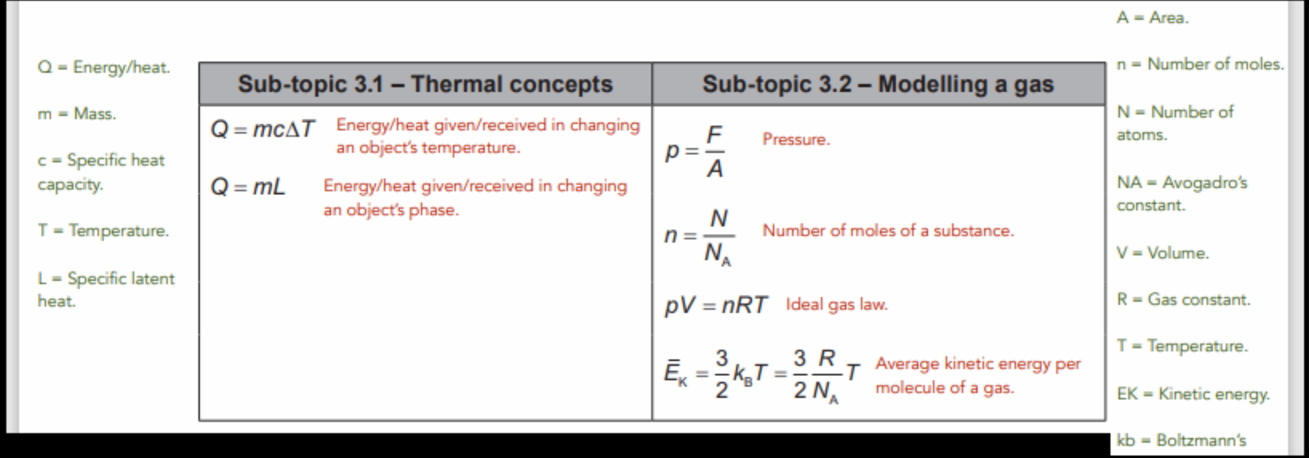
Equating the two gives:

$$1785 + 0.025 \times L_F = 10395 \Rightarrow L \approx 344\,\mathrm{kJ\,kg^{-1}}$$

Nature of science

Models change

As already mentioned, heat was once thought to be a fluid (caloric). Conservation of energy was a natural consequence of this model of heat: a body lost a certain amount of fluid and another gained it. Energy was conserved. So the concept of heat as a fluid seemed natural. But there are phenomena that cannot be explained with this simple picture. For one thing, if heat is a fluid it must have mass. So when heat leaves a body, the



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