## Name of Candidate:

## 11 Temperature

A link between temperature and the behaviour of gas molecules was introduced in Topic 10. In this topic, the concept of temperature is explored in further detail.
Reference to two types of practical thermometer enable aspects of the measurement of temperature to be considered.

## Learning outcomes

Candidates should be able to:
\(\left.$$
\begin{array}{ll}\hline \begin{array}{l}\text { 11.1 Thermal } \\
\text { equilibrium }\end{array} & \text { a) } \begin{array}{l}\text { appreciate that (thermal) energy is transferred from a region of } \\
\text { higher temperature to a region of lower temperature } \\
\text { understand that regions of equal temperature are in thermal } \\
\text { equilibrium }\end{array}
$$ <br>
11.2 Temperature scales \& a) understand that a physical property that varies with temperature <br>
may be used for the measurement of temperature and state <br>

examples of such properties\end{array}\right]\)| b) understand that there is an absolute scale of temperature that does |
| :--- |
| not dependon the property of any particular substance (i.e. the |
| thermodynamic scale and the concept of absolute zero) |
| convert temperatures measured in kelvin to degrees Celsius and |
| recall that $T / \mathrm{K}=T /{ }^{\circ} \mathrm{C}+273.15$ |

## 12 Thermal properties of materials

A simple kinetic model of matter is used to study properties of the three states of matter, including melting and vaporisation.
This topic then introduces the concept of internal energy and the first law of thermodynamics.

Learning outcomes
Candidates should be able to:
12.1 Specific heat capacity and specific latent heat
a) explain using a simple kinetic model for matter:

- the structure of solids, liquids and gases
- why melting and boiling take place without a change in temperature
- why the specific latent heat of vaporisation is higher than specific latent heat of fusion for the same substance
- why a cooling effect accompanies evaporation
define and use the concept of specific heat capacity, and identify the main principles of its determination by electrical methods
c) define and use the concept of specific latent heat, and identify the main principles of its determination by electrical methods
understand that internal energy is determined by the state of the system and that it can be expressed as the sum of a random distribution of kinetic and potential energies associated with the molecules of a system
b) relate a rise in temperature of a body to an increase in its internal energy
c) recall and use the first law of thermodynamics $\Delta U=q+w$ expressed in terms of the increase in internal energy, the heating of the system (energy transferred to the system by heating) and the work done on the system


## Changes of state

- The kinetic model of matter can be used to describe the structure of solids, liquids and gases.
- You should recall that the kinetic model describes the behaviour of matter in terms of moving particles (atoms, molecules, etc.).
- Figure should remind you of how we picture the three states of matter at the
 atomic scale:
- In a solid, the particles are close together, tightly bonded to their
 neighbours, and vibrating about fixed positions.
- In a gas, the particles have broken free from their neighbours; they are widely separated and are free to move around within their container.


## Energy changes:

- Energy must be supplied to raise the temperature of a solid, to melt it, to heat the liquid and to boil it.
- Where does this energy go? It is worth taking a close look at a single change of state and thinking about what is happening on the atomic scale.
- Figure a shows a suitable arrangement.
- A test tube containing octadecanoic acid (a white, waxy substance at room temperature) is warmed in a water bath.
- At $80^{\circ} \mathrm{C}$, the substance
 is a clear liquid.
- The tube is then placed in a rack and allowed to cool.
- Its temperature is monitored, either with a thermometer or with a temperature probe and datalogger.
- Figure b shows typical results.
- The temperature drops rapidly at first, then more
- slowly as it approaches room temperature.
- The important section of the graph is the region $B C$.
- The temperature remains steady for some time.
- The clear liquid is gradually returning to its white, waxy solid state.
- It is essential to note that energy is still being lost even though the temperature is not decreasing.
- When no liquid remains, the temperature starts to drop again.
- From the graph, we can deduce the melting point of octadecanoic acid.
- This is a technique used to help identify substances by finding their melting points


## Heating ice:

- In some ways, it is easier to think of the experiment above in reverse.
- What happens when we heat a substance?
- Imagine taking some ice from the deep freeze.
- Put the ice in a well-insulated container and heat it at a steady rate.
- Its temperature will rise; eventually we will have a container of water vapour. (Note that water vapour is an invisible gas; the 'steam' that you see when a kettle boils is not a gas but a cloud oftiny droplets of liquid water.)
- Figure shows the results we might expect if we could carry out this idealised experiment.
- Energy is supplied to the ice at a constant rate.
- We will consider the
- different sections of this graph in some detail, in order to describe where the energy is going at each stage.
- We need to think about the kinetic and potential energies of the
 molecules.
- If they move around more freely and faster, their kinetic energy has increased.
- If they break free of their neighbours and become more disordered, their electrical potential energy has increased.
- You know that the kinetic energy of a particle is the energy it has due to its motion.
- Figure shows how the electrical potential energy of two isolated atoms depends on their separation.
- Work must be done (energy mustbe put in) to separate neighbouring atoms - think about the work you must do to snap a piece of plastic or to tear a sheet of paper.
- The graph shows that:
- the electrical potential energy of two atoms very close together is large and negative
- as the separation of the atoms increases, their potential
 energy also increases
- when the atoms are completely separated, their potential energy is maximum and has a value of zero.

Now consider the figure again.
Section AB

- The ice starts below $0^{\circ} \mathrm{C}$; its temperature rises.
- The molecules gain energy and vibrate more and more.
- Their vibrational kinetic energy is increasing.
- There is very little change in the mean separation between the molecules and hence very little change in their electrical potential energy.

Section BC

- The ice melts at $0^{\circ} \mathrm{C}$.
- The molecules become more disordered.
- There is a modest increase in their electrical potential energy.


## Section CD

- The ice has become water
- Its temperature rises towards $100^{\circ} \mathrm{C}$.
- The molecules move increasingly rapidly.
- Their kinetic energy is increasing.
- There is very little change in the mean separation between the molecules and therefore very little change in their electrical potential energy.


## Section DE

- The water is boiling.
- The molecules are becoming completely separate from one another.
- There is a large increase in the separation between the molecules and hence their electrical potential energy has increased greatly.
- Their movement becomes very disorderly.


## Section EF

- The steam is being heated above $100^{\circ} \mathrm{C}$.
- The molecules move even faster.
- Their kinetic energy is increasing.
- The molecules have maximum electrical potential energy of zero.
- You should see that, when water is heated, each change of state (melting, boiling) involves the following:
- there must be an input of energy
- the temperature does not change
- the molecules are breaking free of one another
- their potential energy is increasing.

In between the changes of state:

- the input of energy raises the temperature of the substance
- the molecules move faster
o theirkinetic energy is increasing.
- The hardest point to appreciate is that you can put energy into the system without its temperature rising.
- This happens during any change of state; the energy goes to breaking the bonds between neighbouring molecules.
- The energy which must be supplied to cause a change of state is sometimes called 'latent heat'.
- The word 'latent' means 'hidden' and refers to the fact that, when you melt
- something, its temperature does not rise and the energy that you have put in seems to have disappeared.
- It may help to think of temperature as a measure of the average kinetic energy of the molecules.
- When you put a thermometer in some water to measure its temperature,
- the water molecules collide with the thermometer and share their kinetic energy with it.
- At a change of state, there is no change in kinetic energy, so there is no change
- in temperature.
- Notice that melting the ice (section BC) takes much less energy than boiling the same amount of water (section DE).
- This is because, when a solid melts, the molecules are still bonded to most of their immediate neighbours.
- When a liquid boils, each molecule breaks free of all of its neighbours.
- Melting may involve the breaking of one or two bonds per molecule, whereas boiling involves breaking eight or nine.


## Evaporation:

- A liquid does not have to boil to change into a gas.
- A puddle of rain-water dries up without having to be heated to $100^{\circ} \mathrm{C}$.
- When a liquid changes to a gas without boiling, we call this evaporation.
- Any liquid has some vapour associated with it.
- If we think about the microscopic picture of this, we can see why (Figure) Within the liquid, molecules are moving about.
- Some move faster than others, and can break free from the bulk of the liquid. They form the vapour above the liquid.
- Some molecules from the vapour may come back into contact with the surface of
 the liquid, and return to the liquid.
- However, there is a net outflow of energetic molecules from the liquid, and eventually it will evaporate away completely.
- You may have had your skin swabbed with alcohol or ether before an injection.
- You will have noticed how cold your skin becomes as the volatile liquid evaporates.
- Similarly, you can become very cold if you get wet and
- stand around in a windy place.
- This cooling of a liquid is a very important aspect of evaporation.
- When a liquid evaporates, it is the most energetic molecules that are most likely to escape.
- This leaves molecules with a below-average kinetic energy.
- Since temperature is a measure of the average kinetic energy of the molecules, it follows that the temperature of the evaporating liquid must fall.


## Internal energy:

- All matter is made up of particles, which we will refer to here as 'molecules'. Matter can have energy.
- For example, if we lift up a stone, it has gravitational potential energy.
- If we throw it, it has kinetic energy.
- Kinetic and potential energies are the two general forms of energy.
- We consider the stone's potential and kinetic energies to be properties or attributes of the stone itself; we calculate their values ( $m g h$ and $1 / 2 m v^{2}$ ) using the mass and speed of the stone.
- Now think about another way in which we could increase the energy of the stone: we could heat it.
- Now where does the energy from the heater go?
- The stone's gravitational potential and kinetic energies do not increase; it is not higher or faster than before.
- The energy seems to have disappeared into the stone.
- Of course, you already know the answer to this.
- The stone gets hotter, and that means that the molecules which make up the stone have more energy, both kinetic and electrical potential.
- They vibrate more and faster, and they move a little further apart.
- This energy of the molecules is knownas the internal energy of the stone.
- The internal energy of a system (e.g. the heated stone) is defined as follows:
"The internal energy of a system is the sum of the random distribution of kinetic and potential energies of its atoms or molecules."


## Molecular energy:

- Earlier in this chapter, where we studied the phases of matter, we saw how solids, liquids and gases could be characterised by differences in the arrangement, order and motion of their molecules.
- We could equally have said that, in the three phases, the molecules have different amounts of kinetic and potential energy.
- Now, it is a simple problem to find the internal energy of an amount of matter.
- We add up the kinetic and potential
 energies associated with all the molecules in that matter. For example, consider the gas shown in Figure.
- There are ten molecules in the box, each having kinetic and potential energy.
- We can work out what all of these are and add them together, to get the total internal energy of the gas in the box.


## Changing internal energy:

- There are two obvious ways in which we can increase the internal energy of some gas: we can heat it (Figure a), or we can do work on it by compressing it (Figure b).
- Heating a gas
- The walls of the container become hot and so its molecules vibrate more vigorously.
- The molecules of the cool gas strike the walls and bounce off faster.
- They have gained kinetic energy, and we say the temperature has risen.
- Doing work on agas
- In this case, a wall of the container is being pushed inwards.
- The molecules of the cool gas strike a moving wall and bounce off faster.
- They have gained kinetic energy and again the temperature has risen.
- This explains why a gas gets hotter when it is compressed.
- There are other ways in which the internal energy of a system can be increased: by passing an electric current through it, for example.
- However, doing work and heating are all we need to consider here.
- The internal energy of a gas can also decrease for example, if it loses heat to its surroundings, or if it expands so that it does work on its surroundings.


## First law of thermodynamics:

- You will be familiar with the idea that energy is conserved; that is, energy cannot simply disappear, or appear from nowhere.
- This means that, for example, all the energy put into a gas by heating it and by doing work on it must end up in the gas; it increases the internal energy of the gas.

We can write this as an equation:

| increase in <br> internal energy | $=$energy supplied <br> by heating | +Energy supplied by <br> doing work |
| :--- | :---: | :---: | :---: |

In symbols:

$$
\Delta U=q+w
$$

- This is known as the first Jaw of thermo dynamics and is a formal statement of the principle of conservation of energy.
- (It applies to all situations, not simply to a mass of gas.)
- Since you have learned previously that energy is conserved, it may seem to be a simple idea, but it took sçientists a good many decades to understand the nature of energy and to appreciate that it is conserved.


## The meaning of temperature:

- Picture a beaker of boiling water.
- You want to measure its temperature, so you pick up a thermometer which is lying on the bench.
- The thermometer reads $20^{\circ} \mathrm{C}$.
- You place the thermometer in the water and the reading goes up . $30^{\circ} \mathrm{C}$, $40^{\circ} \mathrm{C}, 50^{\circ} \mathrm{C}$.
- This tells you that the thermometer is getting hotter; energy is being transferred from the water to the thermometer.
- Eventually, the thermometer reading reaches $100^{\circ} \mathrm{C}$ and it stops rising.
- Because the reading is steady, you can deduce that energy is no longer being transferred to the thermometer and so its scale tells you the temperature of the water.
- This simple, everyday activity illustrates several points:
- We are used to the idea that a thermometer shows the temperature of something with which it is in contact.
- In fact, it tells you its own temperature.
- As the reading on the scale was rising, it wasn't showing the temperature of the water.
- It was showing that the temperature of the thermometer was rising.
- Energy is transferred from a hotter object to a cooler one.
- The temperature of the water was greater than the temperature of the thermometer, so energy transferred from one to the other.
- When two objects are at the same temperature, there is no transfer of energy between them.

- That is what happened when the thermometer reached the same temperature as the water, so it was safe to say that the reading on the thermometer was the same as the temperature of the water.
- From this, you can see that temperature tells us about the direction in which energy flows.
- If two objects are placed in contact (so that energy can flow between them),
- it will flow from the hotter to the cooler.
- Energy flowing from a region of higher temperature to a region of lower temperature is called thermal energy. (Here, we are not concerned with the mechanism by which the energy is transferred. It may be by conduction, convection or radiation.)
- When two objects, in contact with each other, are at the same temperature, there will be no net transfer of thermal energy between them.
- We say that they are in thermal equilibrium with each other.
- See figure. a Thermal energy is transferred from the hot water to the cooler thermometer because of the temperature difference between them.
- See figure. b When they are at the same temperature, there is no transfer of thermal energy and they are in thermal equilibrium.



## The thermodynamic (Kelvin) scale:

- The Celsius scale of temperature is a familiar, everyday scale of temperature. It is based on the properties of water.
- It takes two fixed points, the melting point of pure ice and the boiling point of pure water, and divides the range between them into 100 equal intervals.
- There is nothing special about these two fixed points.
- In fact, both change if the pressure changes or if the water is impure.
- The thermodynamic scale, also known as the Kelvin scale, is a better scale in that one of its fixed points, absolute zero, has a greater significance than either of the Celsius fixed points.
- It is not possible to have a temperature lower than 0 K .
- Sometimes it is suggested that, at this temperature, matter has no energy left in it .
- This is not strictly trues, it is more correct to say that, for any matter at absolutezero, it is impossible to remove any more energy from it.
- Hence absolute zero is the temperature at which all substances have the minimum internal energy. (The kinetic energy of the atoms or molecules is zero and their electrical potential energy is minimum.)
- We use different symbols to represent temperatures on these two scales: $\theta$ for the Celsius scale, and (T) for the thermodynamic (Kelvin) scale.
- To convert between the two scales, we use these relationships:

$$
\begin{aligned}
& \theta\left({ }^{\circ} \mathrm{C}\right)=\mathrm{T}(\mathrm{~K})-273.15 \\
& \mathrm{~T}(\mathrm{~K})=\theta\left({ }^{\circ} \mathrm{C}\right)+273.15
\end{aligned}
$$

- For most practical purposes, we round off the
 conversion factor to 273 as shown in the conversion chart.
- The thermodynamic scale of temperature is designed to overcome a problem with scales of temperature such as the Celsius scale, which depends on the melting point and boiling point of pure water.
- To measure a temperature on this scale, you might use a liquid-in-glass thermometer.
- However, the expansion of a liquid may be non-linear.
- This means that if you compare the readings from two different types of liquid-in-glass thermometer, for example a mercury thermometer and an alcohol thermometer, you can only be sure that they will agree at the two fixed points
- on the Celsius scale.
- At other temperatures, their readings may differ.
- The thermodynamic scale is said to be an absolute scale as it is not defined in terms of a property of any particular substance.
- It is based on the idea that the average kinetic energy of the particles of a substance increases with temperature.
- The average kinetic energy is the same for all substances at a particular thermodynamic temperature; it does not depend on the material itself.
- In fact, as you will see in upcoming chapter, the average kinetic energy of a gas molecule is proportional to the thermodynamic temperature.
- So, if we can measure the average kinetic energy of the particles of a substance, we can deduce the temperature of that substance.
- The thermodynamic scale has two fixed points:
- absolute zero, which is defined as OK
- the triple point of water, the temperature at which ice, water and water vapour can co-exist, which is defined as 273.16 K (equal to $0.01^{\circ} \mathrm{C}$ ).
- So the gap between absolute zero and the triple point of water is divided into 273.16 equal divisions.
- Each division is 1 K .
- The scale is defined in this slightly odd way so that the scale divisions on the thermodynamic scale are equal in size to the divisions on the Celsius scale, making conversions between the two scales relatively easy.
- A change in temperature of 1 K is thus equal to a change in temperature of $1^{\circ} \mathrm{C}$.


## Thermometers:

- A thermometer is any device which can be used to measure temperature.
- Each type of thermometer makes use of some physical property of a material which changes with temperature.
- The most familiar is the length of a column of liquid in a tube, which gets longer as the temperature increases because the liquid expands - this is how a liquid-in-glass thermometer works.
- Other properties which can be used as the basis of thermometers include:
- the resistance of an electrical resistor or thermistor
- the voltage produced by a thermocouple
- the colour of an electrically heated wire
- the volume of a fixed mass of gas at constant pressure.
- In each case, the thermometer must be calibrated at two or more known temperatures (such as the melting and boiling points of water, which correspond to $0^{\circ} \mathrm{C}$ and $100^{\circ} \mathrm{C}$ ), and the scale between divided into equal divisions.
- There is no guarantee that two thermometers will agree with each other except at these fixed points.
- Now we will look in detail at two types of electrical thermometer.
- In previous chapters, we saw that electrieal resistance changes with temperature.
- For metals, resistance increases with temperature at a fairly steady rate. However, for a thermistor, the resistance changes rapidly over a relatively narrow range of temperatures.
- A small change in temperature results in a large change in resistance, so a thermometer based on a thermistor will be sensitive over that range of temperatures.
- A thermocouple is another electrical device which can be used as the sensor of a thermometer.
- Figure shows the principle.
- Wires of two different metals, $X$ and $Y$, are required.
- A length of metal $X$ has a length of metal
- $Y$ soldered to it at each end. This produces two junctions, which are the important parts of the thermocouple.
- If the two junctions are at different temperatures, an e.m.f. will be produced between the two free ends of the thermocouple, and can be measured using a voltmeter.
- The greater the difference in temperatures, the greater the
 voltage produced; however, this e.m.f. may not vary linearly with temperature, i.e. a graph of e.m.f. against temperature is not usually a straight line.
- Electrical thermometers can measure across areat range of temperatures, from OK to hundreds or even thousands of kelvin.

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Table below compares resistance and thermocouple thermometers.:

| Feature | Resistance thermometer | Thermocouple thermometer |
| :---: | :---: | :---: |
| robustness | very robust | robust |
| range | thermistor: narrow range resistance wire: wide range | can be very wide |
| size | larger than thermocouple, has greater thermal capacity therefore slower acting | smaller than resistance thermometers, has smaller thermal capacity, so quicker acting and can measure temperature at a point |
|  | thermistor: high sensitivity over narrow Vange <br> resistance wire: less sensitive | can be sensitive if appropriate metals chosen |
| linearity | the mister: airly linear over narrows range resistance wire: good linearity | non-linear so requires calibration |
| remote operation | long conducting wires allow the operator be at a distance from the thermometer | long conducting wires allow the operator to be ar a dstance from the thermo meter |

## Calculating energy changes:

- So far, we have considered the effects of heating a substance in qualitative terms, and we have given an explanation in terms of a kinetic model of matter.
- Now we will look at the amount of energy needed to change the temperature of something, and to produce a change of state.
- Specific heat capacity
- If we heat some material so that its temperature rises, the amount of energy we must supply depends on three things:
- the mass $m$ of the material we are heating
- the temperature change $\Delta \theta$ we wish to achieve
- the material itself.
- Some materials are easier to heat than others.
- It takes more energy to raise the temperature of 1 kg of water by $1^{\circ} \mathrm{C}$ than to raise the temperature of 1 kg of alcohol by the same amount.
- We can represent this in an equation.
- The amount of energy E that must be supplied is given by:

$$
E=m c \Delta \theta
$$

- where $c$ is the specific heat capaeity of the material.
- Rearranging this equation gives:

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

- The specific heat capacity of a material can be defined as a word equation as follows:

$$
\text { specific heat capacity }=\frac{\text { energy supplied }}{\text { mass } \times \text { temperature change }}
$$

- Alternatively, specific heat capacity can be defined in words as follows: The specific heat capacity of a substance is the energy required per unit mass of the substance to raise the temperature by 1 K (or $1^{\circ} \mathrm{C}$ ).
- The word 'specific' here means 'per unit mass', i.e. per kg.
- From this form of the equation, you should be able to see that the units of $c$ are $\mathrm{Jkg}^{-1} \mathrm{~K}^{-1}$ (or $\mathrm{Jkg}^{-1}{ }^{\circ} \mathrm{C}^{-1}$ ).

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Table shows some values of specific heat capacity measured at $0^{\circ} \mathrm{C}$.

| Substance | $\mathbf{c} / \mathbf{J ~ k g}^{\mathbf{- 1}} \mathbf{K}^{\mathbf{- 1}}$ |
| :--- | :---: |
| aluminium | 880 |
| copper | 380 |
| lead | 126 |
| glass | $500-680$ |
| ice | 2100 |
| water | 4180 |
| seawater | 3950 |
| ethanol | 2500 |
| mercury | 140 |

- Specific heat capacity is related to the gradient of the sloping sections of the graph shown earlier.
- The steeper the gradient, the faster the substance heatsv up, and hence the lower its specific heat capacity must be.


## How to calculate the specific heat capacity of a substance:

- The metal is in the form of a cylindrical block of mass 1.00 kg .
- An electrical heater is used to supply the energy.
- This type of heater is used because we can easily determine the amount of energy supplied - more easily than if we heated the metal with a Bunsen flame, for example.
- An ammeter and voltmeter are used to make the necessary measurements.
- A thermometer or temperature sensor is used to monitor the block's temperature as it is heated.
- The block must not be heated too quickly; we want to be sure that the energy has time to spread throughout the metal.
- The block should be insulated by wrapping it in a suitable material - this is not shown in the illustration.
- It would be possible in principle to determine $c$ by making justs one measurement of temperature change, but it is better to record values of the temperature as it rises and plot a graph of temperature $\vartheta$ against time $t$.

Sources of error

- This experiment can give reasonably good measurements of specific heat capacities.
- As noted earlier, it is desirable to have a relatively low rate of heating, so that energy spreads throughout the block.
- If the block is heated rapidly, different parts may be at different temperatures.
- Thermal insulation of the material is also vital.
- Inevitably, some energy will escape to the surroundings.
- This means that more energy must be supplied to the block for each degree rise in temperature and so the experimental value for the specific heat capacity will be too high.
- One way around this is to cool the block below room temperature before beginning to heat it.
- Then, as its temperature rises past room temperature, heat losses will be zero in principle, because there is no temperature difference between the block and its surroundings.


## Specific latent heat:

- Energy must be supplied to melt or boil a substance. (In this case, there is no temperature rise to consider since the temperature stays constant during a change of state.)
- This energy is called latent heat.
- The specific latent heat of a substance is the energy required per kilogram of the substance to change its state without any change in temperature.
- When a substance melts, this quantity is called the specific latent heat of fusion; for boiling, it is the specific latent heat of vaporisation.
- To calculate the amount of energy $E$ required to melt or vaporise a mass $m$ of a substance, we simply need to know its specific latent heat $L$ :

$$
E=m L
$$

- $L$ is measured in $\mathrm{Jg}^{-1}$. (Note that there is no 'per ${ }^{\circ} \mathrm{C}^{\prime}$ since there is no change in temperature.)
- For water the values are:
- specific latent heat of fusion of water, $330 \mathrm{~kJ} \mathrm{~kg}-1$
- specific latent heat of vaporisation of water, $2.26 \mathrm{MJ} \mathrm{kg}-1$
- You can see that $L$ for boiling water to form steam is roughly seven times the value for melting ice to form water.
- This is because, when ice melts, only one or two bonds are broken for each molecule; when water boils, several bonds are broken per molecule.
- Determining specific latent heat $L$
- The principle of determining the specific latent heat of a material is similar to determining the specific heat capacity (but remember that there is no change in temperature).
- Figure shows how to measure the specific latent heat of vaporisation of water.
- A beaker containing water is heated using an electrical heater.
- A wattmeter (or an ammeter and a voltmeter) determines the rate at which energy is supplied to the heater.
- The beaker is insulated to minimise energy loss, and it stands on a balance.
- A thermometer is included to ensure that th temperature of the water remains at $100^{\circ} \mathrm{C}$.
- The water is heated at a steady rate and its mass recorded at equal
 intervals of time. Its mass decreases as it boils.
- A graph of mass against time should be a straight line whose gradient is the rate of mass loss.
- The wattmeter shows the rate at which energy is supplied to the water via the heater.
- We thus have:

$$
\text { specific latent head }=\frac{\text { rate of supply of energy }}{\text { rate of loss of mass }}
$$

- A similar approach can be used to determine the specific latent heat of fusion of ice. In this case, the ice is heated electrically in a funnel; water runs out of the funnel and is collected in a beaker on a balance.
- As with any experiment, we should consider sources of error in measuring $L$ and their effects on the final result.
- When water is heated to produce steam, some energy may escape to the surroundings so that the measured energy is greater than that supplied to the water.
- This systematic error gives a value of $L$ which is greater than the true value. When ice is melted, energy from the surroundings will conduct into the ice, so that the measured value of $L$ will be an underestimate.

