Chapter 19

Thermal physics

LEARNING INTENTIONS

In this chapter you will learn how to:

- relate a rise in temperature of an object to internal energy, the sum of the random distribution of kinetic and potential energies of the molecules in a system
- recall and use the first law of thermodynamics
- calculate the work done when the volume of a gas changes at constant pressure
- measure temperature using a physical property and state examples of such properties
- use the thermodynamic scale of temperature, and understand that the lowest possible temperature is zero kelvin and that this is known as absolute zero
- relate transfer of (thermal) energy as being due to a difference in temperature and understand thermal equilibrium
- define and use specific heat capacity and specific latent heat, and outline how these quantities can be measured.

BEFORE YOU START

- Write down the boiling point and melting point of water, and the names scientists use to describe *changes of state* (from solid to liquid, and from liquid to solid, and so on).
- List some difference between atoms and molecules. You can treat them both as simply ‘particles’ and not worry about how many atoms a molecule contains.

FROM WATER TO STEAM

When water boils, it changes state – it turns to water vapour. A liquid has become a gas. This is a familiar process, but Figure 19.1 shows a dramatic example of such a change of state. This is a geyser in New Zealand, formed when water is trapped underground where it is in contact with hot rocks. The temperature and pressure of the water build up until it suddenly erupts above the surface to form a tall plume of scalding water and water vapour. What happens underground to cause this effect? Would this be a useful effect to have happen near where you live?

Figure 19.1: At regular intervals of time, the White Lady Geyser, near Rotorua in New Zealand, throws up a plume of water (liquid) and water vapour (gas).

19.1 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, and so on). Figure 19.2 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.

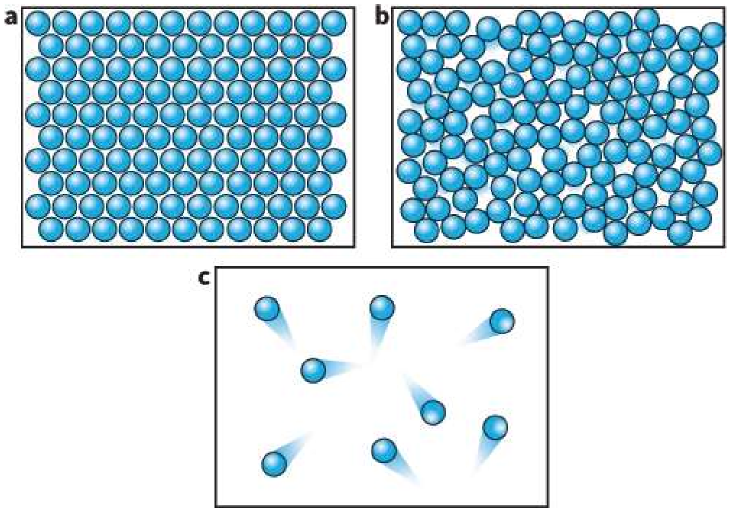

Figure 19.2: Typical arrangements of atoms in **a** a solid, **b** a liquid and **c** a gas.

Question

1 Describe a liquid in terms of the arrangement of its particles, the bonding between them and their motion.

In this chapter, we will extend these ideas to look at the energy changes involved when materials are heated and cooled.

19.2 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 19.3a 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 °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 19.3b shows typical results.

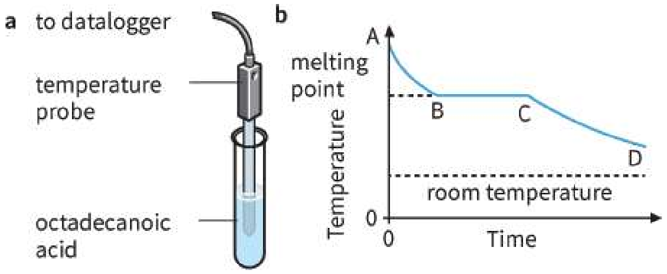

Figure 19.3: **a** Apparatus for obtaining a cooling curve, and **b** typical results.

The temperature drops rapidly at first, then more slowly as it approaches room temperature. The important section of the graph is the region BC. 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 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.

Water vapour and steam mean the same thing–an invisible gas. The ‘steam’ that you see when a kettle boils is *not* a gas; it is ‘wet steam’ – a cloud of tiny droplets of liquid water.

Figure 19.4 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.

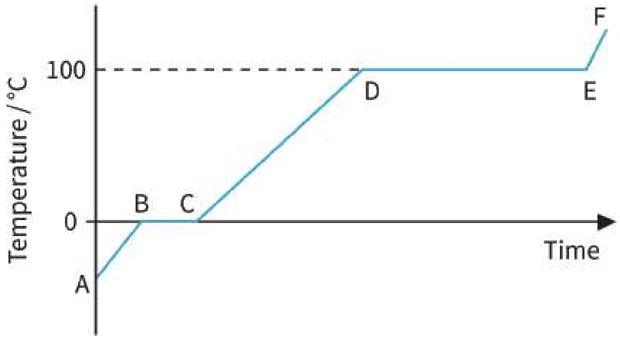

Figure 19.4: A graph of temperature against time for water, heated at a steady rate.

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.

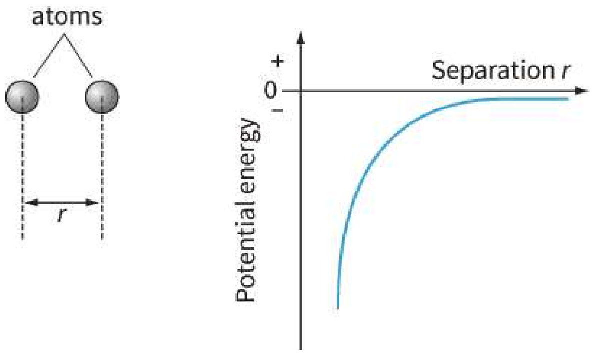

Figure 19.5: The electrical potential energy of atoms is negative and increases as they get further apart.

You know that the kinetic energy of a particle is the energy it has due to its motion. Figure 19.5 shows how the electrical potential energy of two isolated atoms depends on their separation. Work must be done (energy must be 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.

It may seem strange that the potential energy is negative and you will see in Chapter 21 why this is so. At the moment, just notice that, as atoms or molecules become further apart, their potential energy becomes less negative and so they have more potential energy.

Now look back at the graph shown in Figure 19.4.

Section AB

The ice starts below 0 °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 °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 °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 °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
- their kinetic 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 that 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 °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 19.6). 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.

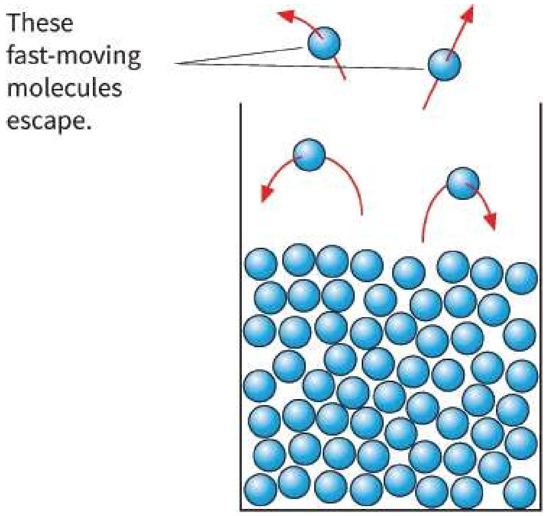

Figure 19.6: Fast-moving molecules leave the surface of a liquid – this is evaporation.

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.

Question

2 Use the kinetic model of matter to explain the following:

a If you leave a pan of water on the hob for a long time, it does not all boil away as soon as the temperature reaches 100 °C.

b It takes less energy to melt a 1.0 kg block of ice at 0 °C than to boil away 1.0 kg of water at 100

°C.

19.3 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 (mgh and $\frac{1}{2}mv^2$) 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 (Figure 19.7). 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.

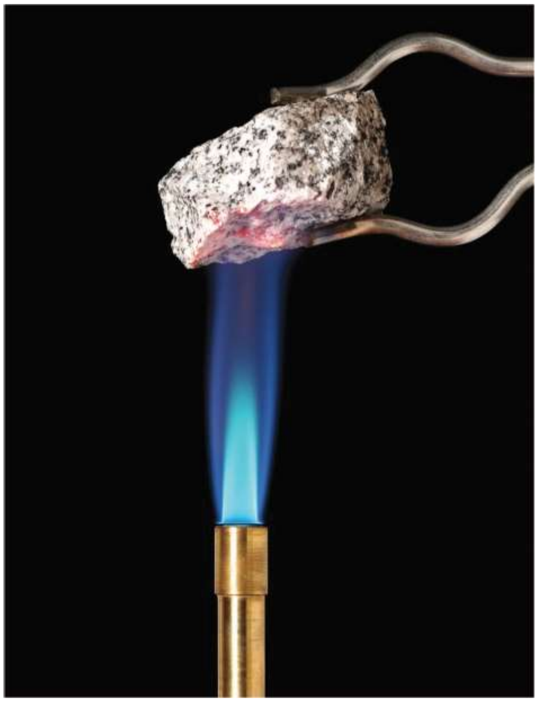

Figure 19.7: Increasing the internal energy of a stone.

Of course, you already know the answer to this. The stone gets hotter, and that means that the molecules that 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 known as the **internal energy** of the stone. The internal energy of a system (such as the heated stone) is defined as 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 19.8. 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.

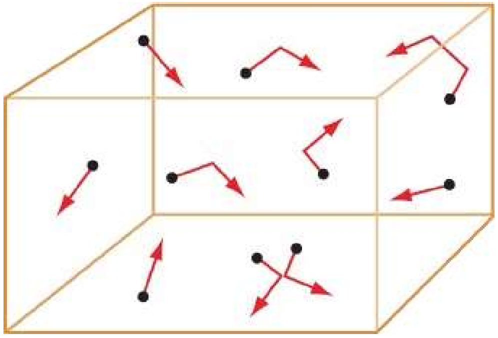

Figure 19.8: The molecules of a gas have both kinetic and potential energy.

Changing internal energy

There are two obvious ways in which we can increase the internal energy of some gas: we can heat it (Figure 19.9a), or we can do work on it by compressing it (Figure 19.9b).

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 a gas

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.

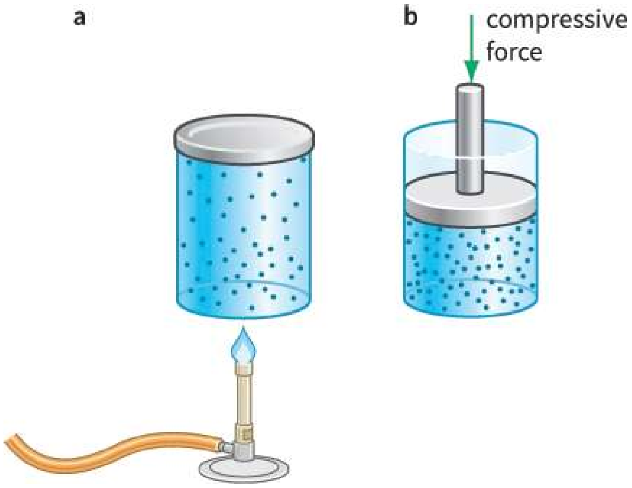

Figure 19.9: Two ways to increase the internal energy of a gas: **a** by heating it, and **b** by compressing it.

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 internal energy = energy supplied by heating + work done **on** the system

In symbols:

$$\Delta U = q + W$$

where ΔU is the increase in internal energy, q is the energy supplied **to** the system by heating and W is the work done on the system

This is known as the **first law of thermodynamics** 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 scientists a good many decades to understand the nature of energy and to appreciate that it is conserved.

You should note the sign convention that is used in the first law. A positive value of ΔU means that the internal energy **increases**, a positive value of q means that heat is added **to** the system, and a positive value of W means that work is done **on** the system. **Negative** values mean that internal energy **decreases**, heat is taken **away** from the system or work is done **by** the system.

Imagine a gas heated from the outside in a sealed container of constant volume. In this case, no work is done on the gas as the heat is added, so W is 0 and the first law equation $\Delta U = q + W$ becomes $\Delta U = q$. All the heat added becomes internal energy of the gas. If the container was able to expand a little as the heat is added then the situation needs some careful thought. Compressing a gas means work is done **on** the gas (W is positive); expanding a gas means work is done **by** the gas (W is negative as it pushes back and does work on the atmosphere). If the container expands then W is slightly negative and ΔU is slightly less than if the volume was constant.

A gas doing work

Gases exert pressure on the walls of their container. If a gas expands, the walls are pushed outwards – the gas has done work on its surroundings (W is negative, if the gas is the system). In a steam engine, expanding steam pushes a piston to turn the engine, and in a car engine, the exploding mixture of fuel and air does the same thing, so this is an important situation.

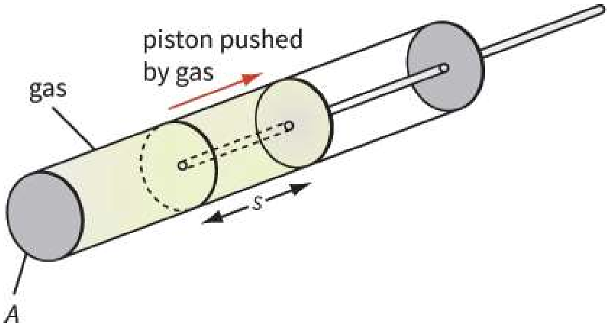

Figure 19.10: When a gas expands, it does work on its surroundings.

Figure 19.10 shows a gas at pressure p inside a cylinder of cross-sectional area A. The cylinder is closed by a moveable piston. The gas pushes the piston a distance s. If we know the force F exerted by the gas on the piston, we can deduce an expression for the amount of work done by the gas.

From the definition of pressure $\left(\text{pressure} = \frac{\text{force}}{\text{area}}\right)$, the force exerted by the gas on the piston is given by:

$$\text{force} = \text{pressure} \times \text{area}$$

$$F = p \times A$$

and the work done is force × displacement:

$$W = p \times A \times s$$

But the quantity $A \times s$ is the **increase** in volume of the gas; that is, the shaded volume in Figure 19.10. We call this ΔV, where the Δ indicates that it is a change in V. Hence, the work done by the gas in expanding is:

$$W = p\Delta V$$

KEY EQUATION

$$W = p\Delta V$$

Work done when the volume of a gas changes at constant pressure.

Notice that we are assuming that the pressure p does not change as the gas expands. This will be true if the gas is expanding against the pressure of the atmosphere, which changes only very slowly.

How does the first law of thermodynamics apply if you compress a gas? This can be done in different ways but we can consider two limiting ways.

Not allowing heat to enter or leave the system

This can be done by pushing the piston into the syringe very fast or by insulating the syringe. In this case, q is zero and $\Delta U = W$. All the work done by pushing in the piston increases the internal energy of the molecules. In this case, the kinetic energy of the molecules increases and the temperature increases, unless there is a change of state.

At constant temperature

Imagine pushing the piston very slowly into a syringe containing gas; so slowly that the temperature stays constant at room temperature. This change is known as an **isothermal change**. The kinetic energy of the molecules remains constant.

The molecules become slightly closer together and this may mean that their internal energy U becomes slightly less but the change is very small (unless the gas becomes a liquid). If U is constant, then ΔU is zero and $0 = q + W$. This means that, if you push the piston in and do positive work W, then q is negative, and heat is lost from the syringe. You can think of this as doing positive work on the system and, with no extra internal energy, the system must lose some heat to the surroundings, perhaps by conduction of heat through the walls of the syringe. Similarly, if you pull the piston out very slowly, W is negative and q is positive and heat enters the system.

Questions

3 Use the first law of thermodynamics to answer the following.

- **a** A gas is heated by supplying it with 250 kJ of thermal energy; at the same time, it is compressed so that 500 kJ of work is done on the gas. Calculate the change in the internal energy of the gas.
- **b** The same gas is heated as before with 250 kJ of energy. This time the gas is allowed to expand so that it does 200 kJ of work on its surroundings. Calculate the change in the internal energy of the gas.

4 When you blow up a balloon, the expanding balloon pushes aside the atmosphere. How much work is done against the atmosphere in blowing up a balloon from a very small volume to a volume of 2 litres ($0.002\ \mathrm{m}^3$)? (Atmospheric pressure = $1.0 \times 10^5\ \mathrm{N\ m}^{-2}$.)

19.4 The meaning of temperature

Picture a beaker of boiling water. You want to measure its temperature, so you pick up a thermometer that is lying on the bench. The thermometer reads 20 °C. You place the thermometer in the water and the reading goes up … 30 °C, 40 °C, 50 °C. This tells you that the thermometer is getting hotter; energy is being transferred from the water to the thermometer.

KEY IDEA

Thermal energy is transferred from a region of higher temperature to a region of lower temperature.

Eventually, the thermometer reading reaches 100 °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 are at the same temperature, they are in **thermal equilibrium** with each other. There will be no net transfer of thermal energy between them when they are in contact with each other – see Figure 19.11.

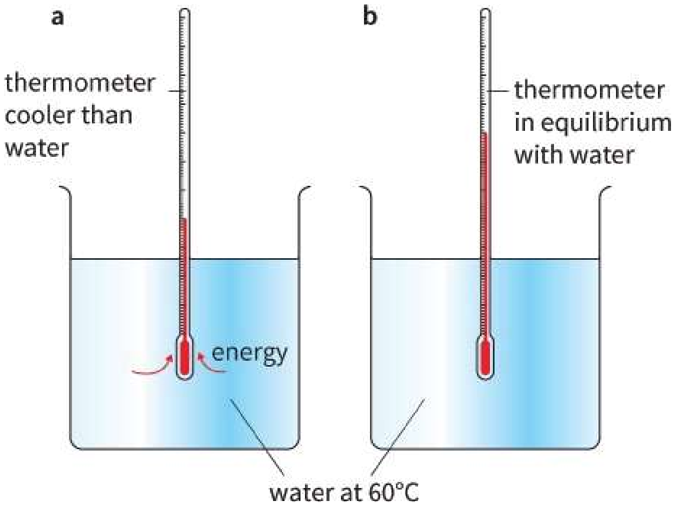

Figure 19.11: **a** Thermal energy is transferred from the hot water to the cooler thermometer because of the temperature difference between them. **b** When they are at the same temperature, there is no transfer of thermal energy and they are in thermal equilibrium.

The thermodynamic (Kelvin) scale

The Celsius scale of temperature is a familiar, everyday scale of temperature. It was originally based on the properties of water with the melting point of pure ice as 0°C and the boiling point of pure water as 100°C.

There is nothing special about these two temperatures. In fact, both the melting point and boiling point 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**, is very important.

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 true; it is more correct to say that, for any matter at absolute zero, 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: θ for the Celsius scale, and T for the thermodynamic (Kelvin) scale. To convert between the two scales, we use these relationships:

$$\theta\,(\text{in}^\circ C) = T\,(\text{in } K) - 273.15$$
$$T\,(\text{in } K) = \theta\,(\text{in }^\circ C) + 273.15$$

For most practical purposes, we round off the conversion factor to 273 as shown in the conversion chart (Figure 19.12).

KEY EQUATIONS

$$\theta\,(\text{in}^\circ C) = T\,(\text{in } K) - 273.15$$
$$T\,(\text{in } K) = \theta\,(\text{in }^\circ C) + 273.15$$

To convert temperatures between degrees Celsius and Kelvin.

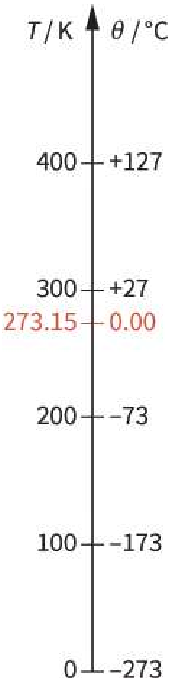

Figure 19.12: A conversion chart relating temperatures on the thermodynamic (Kelvin) and Celsius scales.

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.

KEY IDEA

Thermodynamic temperatures do not depend on the property of any particular substance.

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 Chapter 20, 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 0 K
- 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 °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 °C.

19.5 Thermometers

A thermometer is any device that can be used to measure temperature. Each type of thermometer makes use of some physical property of a material that 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; it depends on a change in density of a liquid. Other physical properties that can be used as the basis of thermometers include:

- the resistance of an electrical resistor or thermistor
- the e.m.f. (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 °C and 100 °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.

Questions

5 **a** Convert each of the following temperatures from the Celsius scale to the thermodynamic scale: 0 °C, 20 °C, 120 °C, 500 °C, −23 °C, −200 °C.

b Convert each of the following temperatures from the thermodynamic scale to the Celsius scale: 0 K, 20 K, 100 K, 300 K, 373 K, 500 K.

6 The electrical resistance of a pure copper wire is mostly due to the vibrations of the copper atoms. Table 19.1 shows how the resistance of a length of copper wire is found to change as it is heated. Copy the table and add a column showing the temperatures in K. Draw a graph to show these data. (Start the temperature scale of your graph at 0 K.) Explain why you might expect the resistance of copper to be zero at this temperature.

Temperature / °C	Resistance / Ω
10	3120
50	3600
75	3900
100	4200
150	4800
220	5640
260	6120

Table 19.1: The variation of resistance with temperature for a length of copper wire.

In Chapter 10, we saw that electrical 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 19.13 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 a great **range** of temperatures, from 0 K to hundreds or even thousands of kelvin.

Table 19.2 compares resistance and thermocouple thermometers.

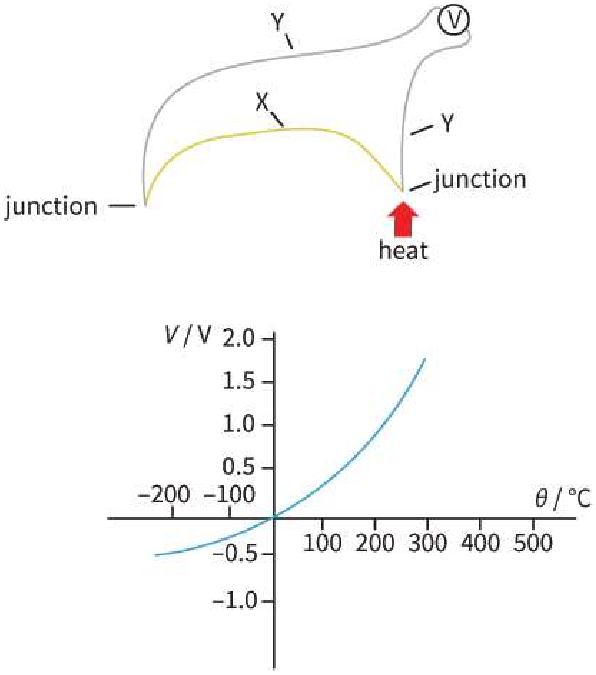

Figure 19.13: The construction of a thermocouple thermometer; the voltage produced depends on the temperature (as shown in the calibration graph) and on the metals chosen.

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
sensitivity	thermistor: high sensitivity over narrow range resistance wire: less sensitive	can be sensitive if appropriate metals chosen
linearity	thermistor: fairly linear over narrow range resistance wire: good linearity	non-linear so requires calibration
remote operation	long conducting wires allow the operator to be at a distance from the thermometer	long conducting wires allow the operator to be at a distance from the thermometer

Table 19.2: Comparing resistance and thermocouple thermometers.

Question

7 Give **one** word for each of the following:

a adding a scale to a thermometer

b all the temperatures, from lowest to highest, which a thermometer can measure

c the extent to which equal rises in temperature give equal changes in the thermometer's output

d how big a change in output is produced by a given change in temperature.

19.6 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
- temperature change Δq we wish to achieve
- 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 °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:

$$E = mc\Delta\theta$$

where c is the **specific heat capacity** 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:

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

Alternatively, specific heat capacity can be defined in words as follows:

The numerical value of 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 °C).

The word ‘specific’ here means ‘per unit mass’; that is, per kg. From this form of the equation, you should be able to see that the units of c are J kg^{-1} K^{-1} (or J kg^{-1} $°\text{C}^{-1}$). Table 19.3 shows some values of specific heat capacity measured at 0 °C.

Specific heat capacity is related to the gradient of the sloping sections of the graph shown earlier in Figure 19.4. The steeper the gradient, the faster the substance heats up and hence the lower its specific heat capacity must be. Worked example 1 shows how to calculate the specific heat capacity of a substance.

WORKED EXAMPLE

1 When 26 400 J of energy is supplied to a 2.0 kg block of aluminium, its temperature rises from 20 °C to 35 °C. The block is well insulated so that there is no energy loss to the surroundings. Determine the specific heat capacity of aluminium.

Step 1 We are going to use the equation:

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

We need to write down the quantities that we know:

E = 26 400 J $\quad$ m = 2.0 kg

$\Delta\theta$ = (35 − 20) °C = 15 °C $\quad$ (or 15 K)

Step 2 Now substitute these values and solve the equation:

$$\begin{aligned} c &= \frac{E}{m\Delta\theta} \\ &= \frac{26\,400}{(2.0 \times 15)} \\ &= 880\ \text{J kg}^{-1}\text{K}^{-1} \end{aligned}$$

Substance	c / J kg^{-1} K^{-1}
aluminium	880
copper	380

copper	380
lead	126
glass	500–680
ice	2100
water	4180
seawater	3950
ethanol	2500
mercury	140

Table 19.3: Values of specific heat capacity.

Questions

You will need to use data from Table 19.3 to answer these questions.

8 Calculate the energy that must be supplied to raise the temperature of 5.0 kg of water from 20 °C to 100 °C.

9 Which requires more energy – heating a 2.0 kg block of lead by 30 K or heating a 4.0 kg block of copper by 5.0 K?

10 A well-insulated 1.2 kg block of iron is heated using a 50 W heater for 4.0 min. The temperature of the block rises from 22 °C to 45 °C. Find the experimental value for the specific heat capacity of iron.

PRACTICAL ACTIVITY 19.1

Determining specific heat capacity *c*

How can we determine the specific heat capacity of a material? The principle is simple: supply a known amount of energy to a known mass of the material and measure the rise in its temperature. Figure 19.14 shows one practical way of doing this for a metal.

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.

Figure 19.14: A practical arrangement for determining the specific heat capacity of a metal.

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 just one measurement of temperature change, but it is better to record values of the temperature as it rises and plot a graph of temperature θ against time t. The method of calculating *c* is illustrated in Worked example 2.

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.

WORKED EXAMPLE

2 An experiment to determine the specific heat capacity c of a 1.00 kg aluminium block is carried out; the block is heated using an electrical heater. The current in the heater is 4.17 A and the p.d. across it is 12 V. Measurements of the rising temperature of the block are represented by the graph shown in Figure 19.15. Determine a value for the specific heat capacity c of aluminium.

Step 1 Write down the equation that relates energy change to specific heat capacity:

$$E = mc\Delta\theta$$

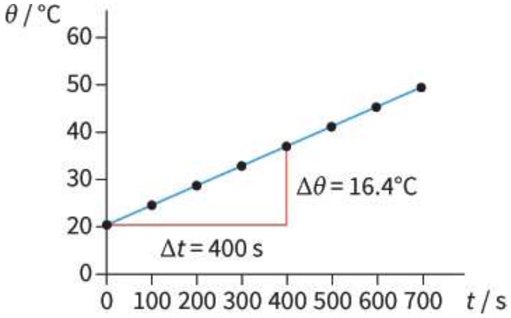

Figure 19.15: Graph of temperature against time for an aluminium block as it is heated.

Step 2 Divide both sides by a time interval Δt:

$$\frac{E}{\Delta t} = mc\left(\frac{\Delta\theta}{\Delta t}\right)$$

The quantity $\frac{E}{\Delta t}$ is the rate at which energy is supplied; that is, the power P of the heater.

The quantity $\frac{\Delta\theta}{\Delta t}$ is the rise of temperature of the block; that is, the gradient of the graph of θ against t.

Hence: $P = m \times c \times$ gradient

Step 3 Calculate the power of the heater and the gradient of the graph.

power = p.d. × current

$P = VI = 12 \times 4.17 \approx 50$ W

$$\begin{aligned} \text{gradient} &= \frac{\Delta\theta}{\Delta t} \\ &= \frac{16.4}{400} \\ &= 0.041\,^{\circ}\mathrm{C s^{-1}} \end{aligned}$$

Step 4 Substitute values, rearrange and solve.

$$\begin{aligned} 50 &= 1.00 \times c \times 0.041 \\ c &= \frac{50}{(1.00 \times 0.041)} \\ c &= 1220\ \mathrm{J\,kg^{-1}K^{-1}} \end{aligned}$$

Questions

11 At higher temperatures than shown, the graph in Figure 19.15 deviates increasingly from a straight line. Suggest an explanation for this.

12 In measurements of the specific heat capacity of a metal, energy losses to the surroundings are a source of error. Is this a systematic error or a random error? Justify your answer.

13 In an experiment to measure the specific heat capacity of water, a student uses an electrical heater to heat some water. His results are shown. Calculate a value for the heat capacity of water. Comment on any likely sources of error.

mass of beaker = 150 g
mass of beaker + water = 672 g
current in the heater = 3.9 A
p.d. across the heater = 11.4 V
initial temperature = 18.5 °C
final temperature = 30.2 °C
time taken = 13.0 min

14 A block of paraffin wax was heated gently, at a steady rate. Heating was continued after the wax had completely melted. The graph of Figure 19.16 shows how the material's temperature varied during the experiment.

a For each section of the graph (AB, BC and CD), describe the state of the material.

b For each section, explain whether the material's internal energy is increasing, decreasing or remaining constant.

c Consider the two sloping sections of the graph. State whether the material's specific heat capacity is greater when it is a solid or when it is a liquid. Justify your answer.

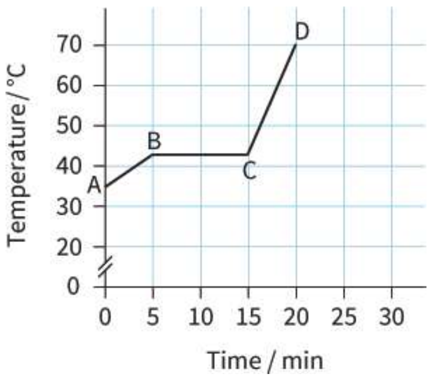

Figure 19.16: Temperature variation of a sample of wax, heated at a constant rate.

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 numerical value of 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:

$$E = mL$$

L is measured in J kg^{-1}. (Note that there is no 'per °C' since there is no change in temperature.) For water, the values are:

- specific latent heat of fusion of water, 330 kJ kg^{-1}
- specific latent heat of vaporisation of water, 2.26 MJ 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. As we saw previously in the topic on heating ice, 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. Worked example 3 shows how to calculate these amounts of energy.

WORKED EXAMPLE

3 The specific latent heat of vaporisation of water is 2.26 MJ kg^{-1}. Calculate the energy needed to change 2.0 g of water into steam at 100 °C.

Step 1 We have been given the following quantities:

$m = 2.0\ \text{g} = 0.002\ \text{kg}$ and $L = 2.26\ \text{MJ kg}^{-1}$

Step 2 Substituting these values in the equation $E = mL$, we have:

$\text{energy} = 0.002 \times 2.26 \times 10^6 = 4520\ \text{J}$

Questions

15 The specific latent heat of fusion of water is 330 kJ kg^{-1}. Calculate the energy needed to change 2.0 g of ice into water at 0 °C. Suggest why the answer is much smaller than the amount of energy calculated in Worked example 3.

16 A sample of alcohol is heated with a 40 W heater until it boils. As it boils, the mass of the liquid decreases at a rate of 2.25 g per minute. Assuming that 80% of the energy supplied by the heater is transferred to the alcohol, estimate the specific latent heat of vaporisation of the alcohol. Give your answer in kJ kg^{-1}.

PRACTICAL ACTIVITY 19.2

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 19.17 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 the temperature of the water remains at 100 °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:

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

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.

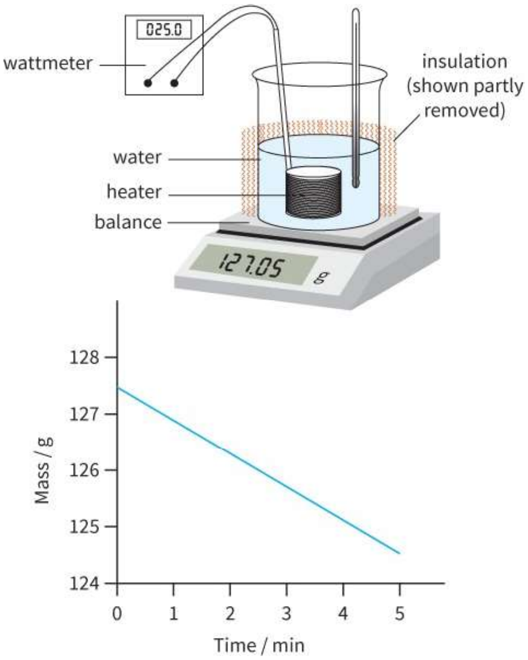

Figure 19.17: Determining the specific latent heat of vaporisation of water.

REFLECTION

List all the ideas in this chapter that are associated with an **increase** in temperature.

What strategies could you use to make sure you understand these?

SUMMARY

The kinetic model of matter allows us to explain behaviour (such as changes of state) and macroscopic properties (such as specific heat capacity and specific latent heat) in terms of the behaviour of molecules.

The internal energy of a system is the sum of the random distribution of kinetic and potential energies associated with the atoms or molecules that make up the system.

If the temperature of an object increases, there is an increase in its internal energy.

Internal energy also increases during a change of state, but there is no change in temperature.

The first law of thermodynamics expresses the conservation of energy:

increase in internal energy = energy supplied by heating + work done on the system

$W = p\Delta V$ is the work done **on** a gas when the volume of a gas changes at constant pressure. W is positive when the gas is compressed (ΔV is negative).

Temperatures on the thermodynamic (Kelvin) and Celsius scales of temperature are related by:

$$T\ (\text{K}) = \theta\ (°\text{C}) + 273.15 \qquad \theta\ (°\text{C}) = T\ (\text{K}) - 273.15$$

Absolute zero, 0 K, is the lowest possible temperature.

A thermometer makes use of a physical property of a material that varies with temperature.

The word equation for the specific heat capacity of a substance is:

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

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 °C).

The energy transferred in raising the temperature of a substance is given by:

$$E = mc\Delta\theta$$

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:

$$E = mL$$

EXAM-STYLE QUESTIONS

1 The first law of thermodynamics can be represented by the expression: $\Delta U = q + W$.

An ideal gas is compressed at constant temperature.

Which row shows whether ΔU, q and W are negative, positive or zero during the change? **[1]**

	ΔU	q	W
A	negative	negative	positive
B	positive	positive	negative
C	zero	negative	positive
D	zero	positive	negative

2 What is the internal energy of an object? **[1]**

A the amount of heat supplied to the object

B the energy associated with the random movement of the atoms in the object

C the energy due to the attraction between the atoms in the object

D the potential and kinetic energy of the object.

3 Describe the changes to the kinetic energy, the potential energy and the total internal energy of the molecules of a block of ice as:

a it melts at 0 °C **[3]**

b the temperature of the water rises from 0 °C to room temperature. **[3]**

[Total: 6]

4 Explain, in terms of kinetic energy, why the temperature of a stone increases when it falls from a cliff and lands on the beach below. **[3]**

5 Explain why the barrel of a bicycle pump gets very hot as it is used to pump up a bicycle tyre. (Hint: the work done against friction is not large enough to explain the rise in temperature.) **[3]**

6 The so-called ‘zeroth law of thermodynamics’ states that if the temperature of body A is equal to the temperature of body B and the temperature of body B is the same as body C, then the temperature of body C equals the temperature of body A.

Explain, in terms of energy flow, why the concept of temperature would be meaningless if this law was not obeyed. **[2]**

7 **a** The first law of thermodynamics can be represented by the expression: $\Delta U = q + W$. State what is meant by all the symbols in this expression. **[3]**

b Figure 19.18 shows a fixed mass of gas that undergoes a change from A to B and then to C.

i During the change from A to B, 220 J of thermal energy (heat) is removed from the gas. Calculate the change in the internal energy of the gas. **[2]**

ii During the change from B to C, the internal energy of the gas decreases by 300 J. Using the first law of thermodynamics explain how this change can occur. **[2]**

[Total: 7]

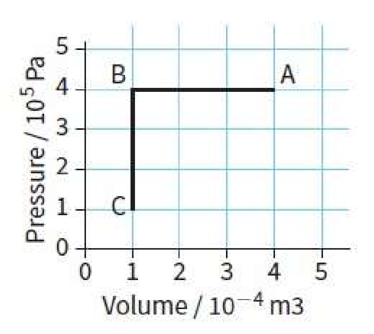

Figure 19.18

8 When a thermocouple has one junction in melting ice and the other junction in boiling water it produces an e.m.f. of 63 µV.

a What e.m.f. would be produced if the second junction was also placed in melting ice? **[1]**

b When the second junction is placed in a cup of coffee, the e.m.f. produced is 49 µV. Calculate the temperature of the coffee. **[2]**

c The second junction is now placed in a beaker of melting lead at 327 °C.

i Calculate the e.m.f. that would be produced. **[2]**

ii State the assumption that you make. **[1]**

[Total: 6]

9 The resistance of a thermistor at °C is 2000 Ω. At 100 °C the resistance falls to 200 Ω. When the thermistor is placed in water of constant temperature, its resistance is 620 Ω.

a Assuming that the resistance of the thermistor varies linearly with temperature, calculate the temperature of the water. **[2]**

b The temperature of the water on the thermodynamic scale is 280 K.

By reference to the particular features of the thermodynamic scale of temperature, **comment** on your answer to part **a**. **[3]**

[Total: 5]

10 a A 500 W kettle contains 300 g of water at 20 °C. Calculate the minimum time it would take to raise the temperature of the water to boiling point. **[5]**

b The kettle is allowed to boil for 2 minutes. Calculate the mass of water that remains in the kettle. State any assumptions that you make. **[4]**

(Specific heat capacity of water = 4.18×10^3 J kg^{-1} °C^{-1}; specific latent heat of vaporisation of water = 2.26×10^6 J kg^{-1}.)

[Total: 9]

11 a Define specific heat capacity of a substance. **[2]**

b A mass of 20 g of ice at −15 °C is taken from a freezer and placed in a beaker containing 200 g of water at 26 °C. Data for ice and water are given in Table 19.5.

	Specific heat capacity / J kg^{-1} K^{-1}	**Specific latent heat of fusion / J kg^{-1}**
ice	2.1×10^3	3.3×10^5
water	4.2×10^3	

Table 19.5

i Calculate the amount of thermal energy (heat) needed to convert all the ice at −15 °C to water at 0 °C. **[2]**

ii Calculate the final temperature *T* of the water in the beaker assuming that the beaker has negligible mass. **[3]**

[Total: 7]

12 a Define specific latent heat and explain the difference between latent heat of fusion and latent heat of vaporisation. **[3]**

b An electric heater generating power of 120 W is immersed in a beaker of liquid that is placed on a balance. When the liquid begins to boil it is noticed that the mass of the beaker and liquid decreases by 6.2 g every minute.

i State how this shows that the liquid is boiling at a steady rate. **[1]**

ii Calculate a value for the specific latent heat of vaporisation of the liquid. **[2]**

iii State and explain whether the value determined in **ii** is likely to be larger or smaller than the accepted value. **[2]**

[Total: 8]

13 a Explain why energy is needed for boiling even though the temperature of the liquid remains constant. **[2]**

This diagram shows an apparatus that can be used to measure the specific latent heat of vaporisation of nitrogen.

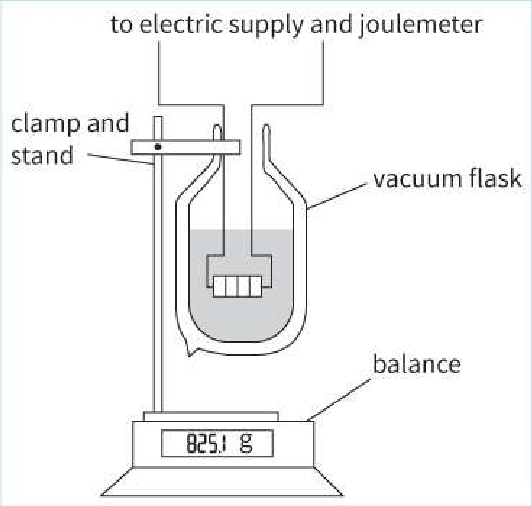

Figure 19.19

b Suggest why the nitrogen is contained in a vacuum flask. **[1]**

c The change in mass of the nitrogen is measured over a specific time interval with the heater switched off. The heater is switched on, transferring energy at 40 W, and the change of mass is found once more.

The results are shown in the table.

	Initial reading on balance / g	Final reading on balance / g	Time / minutes
heater off	834.7	825.5	4
heater on	825.5	797.1	2

Table 19.6

Calculate the specific latent heat of vaporisation of liquid nitrogen. **[4]**

[Total: 7]

14 a i Explain what is meant by internal energy. **[2]**

ii Explain what is meant by the absolute zero of temperature. **[2]**

b An electric hot water heater has a power rating of 9.0 kW. The water is heated as it passes through the heater. Water flows through the heater at a speed of 1.2 m s^{-1} through pipes that have a total cross-sectional area of 4.8×10^{-5} m^2. The temperature of the water entering the heater is 15 °C.

i Calculate the mass of water flowing through the heater each second. **[2]**

ii Calculate the temperature at which the water leaves the heater. **[3]**

iii State any assumptions you have made in your calculation. **[1]**

iv It is possible to adjust the temperature of the water from the heater. Suggest how the temperature of the water could be increased without increasing the power of the heater. **[1]**

(Density of water = 1000 kg m^{-3}; specific heat capacity of water = 4200 J kg^{-1} $°C^{-1}$.)

[Total: 11]

SELF-EVALUATION CHECKLIST

After studying the chapter, complete a table like this:

I can	See topic...	Needs more work	Almost there	Ready to move on
understand internal energy, determined by the state of the system, as the sum of a random distribution of kinetic and potential energies of the molecules of a system	19.3			
relate a rise in temperature of an object to an increase in its internal energy	19.3			
recall and use $W = p\Delta V$ for the work done when the volume of a gas changes at constant pressure and understand the difference between work done by a gas and the work done on a gas	19.3			
recall and use the first law of thermodynamics: $\Delta U = q + W$	19.3			
understand that (thermal) energy is transferred from a region of higher temperature to a region of lower temperature and that regions of equal temperature are in thermal equilibrium	19.4			
understand the use of a physical property that varies with temperature to measure temperature and state examples of such properties	19.5			
understand that thermodynamic temperature does not depend on the property of any particular substance and recall and use: $T / \text{K} = \theta / {}^\circ\text{C} + 273.15$	19.4			
understand that the lowest possible temperature is zero kelvin, absolute zero	19.4			
define and use specific heat capacity	19.6			
define and use specific latent heat and distinguish between specific latent heat of fusion and specific latent heat of vaporisation.	19.6			

Chapter 20

Ideal gases

LEARNING INTENTIONS

In this chapter you will learn how to:

- measure amounts of a substance in moles and find the number of particles using molar quantities
- solve problems using the equation of state $pV = nRT$ for an ideal gas
- deduce a relationship between pressure, volume and the microscopic properties of the molecules of a gas, stating the assumptions of the kinetic theory of gases
- relate the kinetic energy of the molecules of a gas to its temperature and calculate root-mean-square speeds.

BEFORE YOU START

- With a classmate, write down what you know about Brownian motion and what it shows about the molecules in a gas.
- Try to explain to a classmate, in terms of momentum change, why a ball hitting a wall exerts a force on it.
- List Newton’s laws of motion.

THE IDEA OF A GAS

Figure 20.1 shows a weather balloon being launched. Balloons like this carry instruments high into the atmosphere, to measure pressure, temperature, wind speed and other variables.

The balloon is filled with helium so that its overall density is less than that of the surrounding air. The result is an upthrust on the balloon, greater than its weight, so that it rises upwards. As the balloon moves upwards, the pressure of the surrounding atmosphere decreases so that the balloon expands. The temperature drops, which tends to make the gas in the balloon shrink. In this chapter, we will look at the behaviour of gases as their pressure, temperature and volume change.