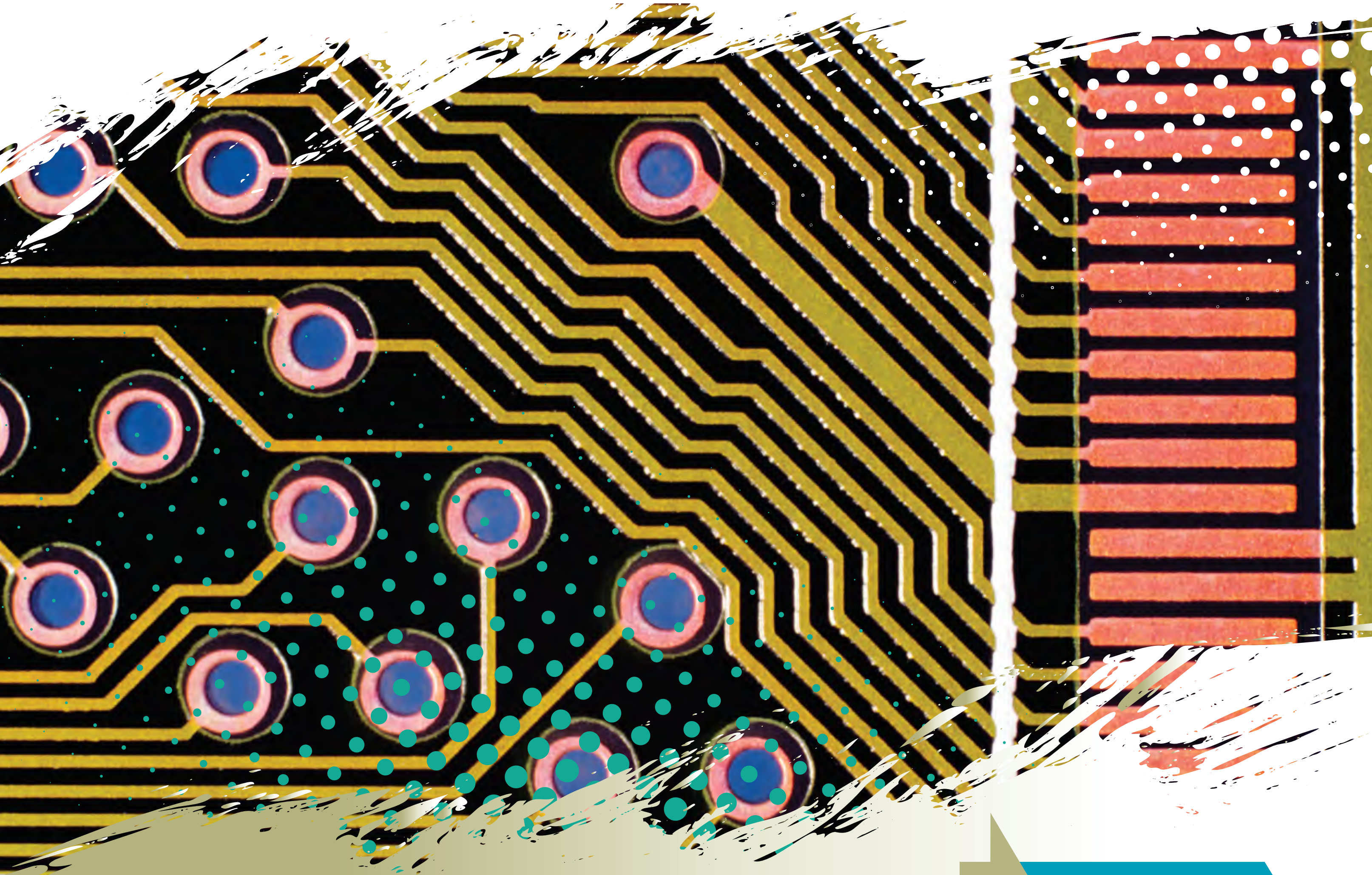
PEARSON

PHYSICS

QUEENSLAND

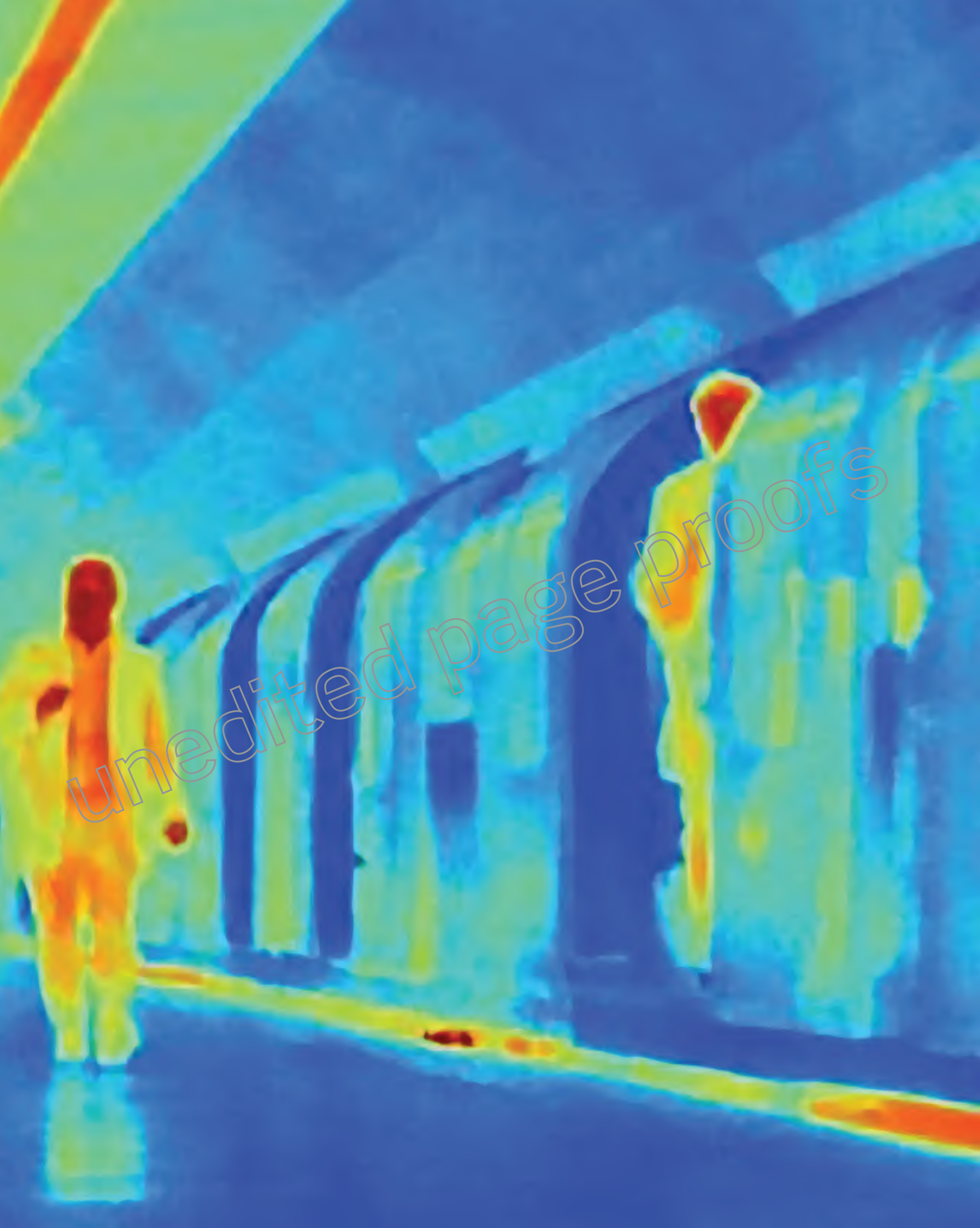
UNITS 1 & 2

2nd EDITION



QCE 2025 Physics SYLLABUS

Student Book



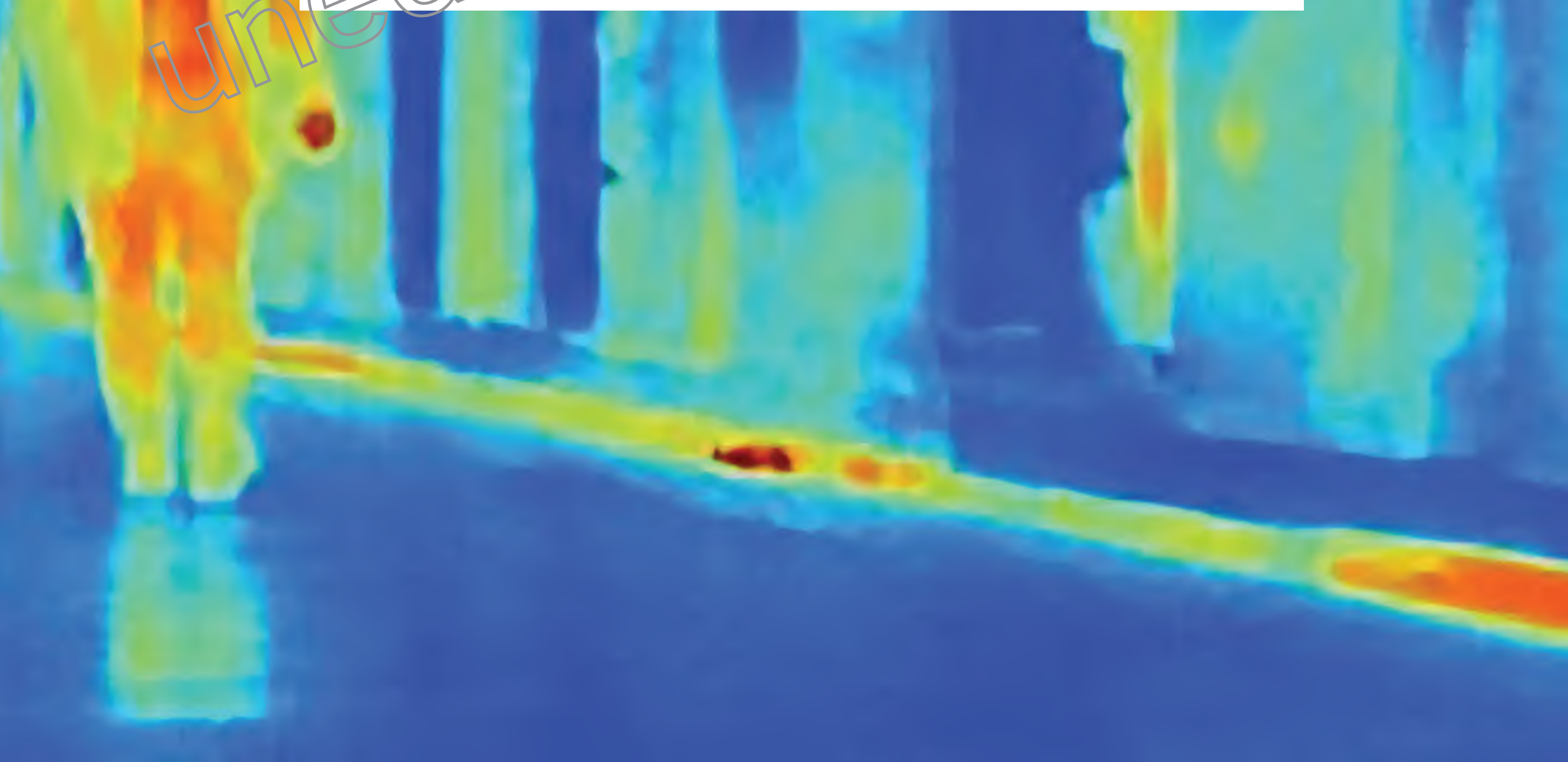
UNIT 1

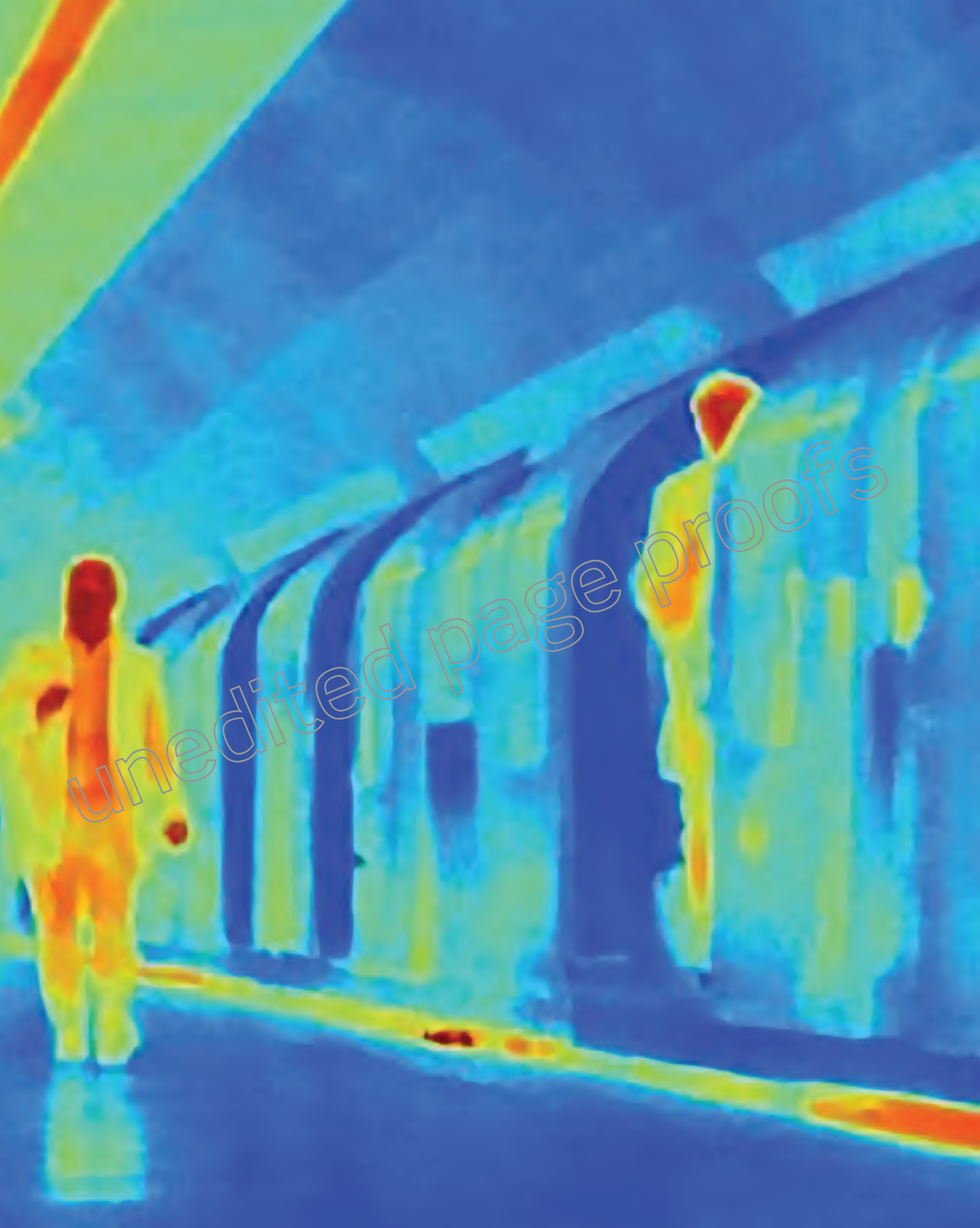
Thermal, nuclear and electrical physics

Unit 1 objectives

Students will:

- describe ideas and findings about heating processes, ionising radiation and nuclear reactions, and electrical circuits
- apply understanding of heating processes, ionising radiation and nuclear reactions, and electrical circuits
- analyse data about heating processes, ionising radiation and nuclear reactions, and electrical circuits
- interpret evidence about heating processes, ionising radiation and nuclear reactions, and electrical circuits
- evaluate processes, claims and conclusions about heating processes, ionising radiation and nuclear reactions, and electrical circuits
- investigate phenomena associated with heating processes, ionising radiation and nuclear reactions, and electrical circuits.





CHAPTER 02

Heating processes

WS 1.1.1

Thermal energy is part of our everyday experience. Humans can thrive in Earth's climatic extremes, from the outback deserts to ski slopes in winter. Due to increasing levels of carbon dioxide in the atmosphere, the Earth is getting warmer. In 2023, the Earth was the warmest since records began in 1890. These higher temperatures increase the severity of bushfires and increase the rate of evaporation from pastures, which dries the land and reduces the level of food production.

Temperature is a measure of the motion of particles. Different substances will absorb different amounts of heat to raise their temperature or to change state from solid to liquid to gas, depending on their atomic structure. Heat energy can be transmitted through materials in different ways, depending on the material's properties. Energy is conserved in the universe, but in each energy transfer, energy can be 'wasted' as less useful forms, decreasing the efficiency of the energy transfer.

Syllabus subject matter

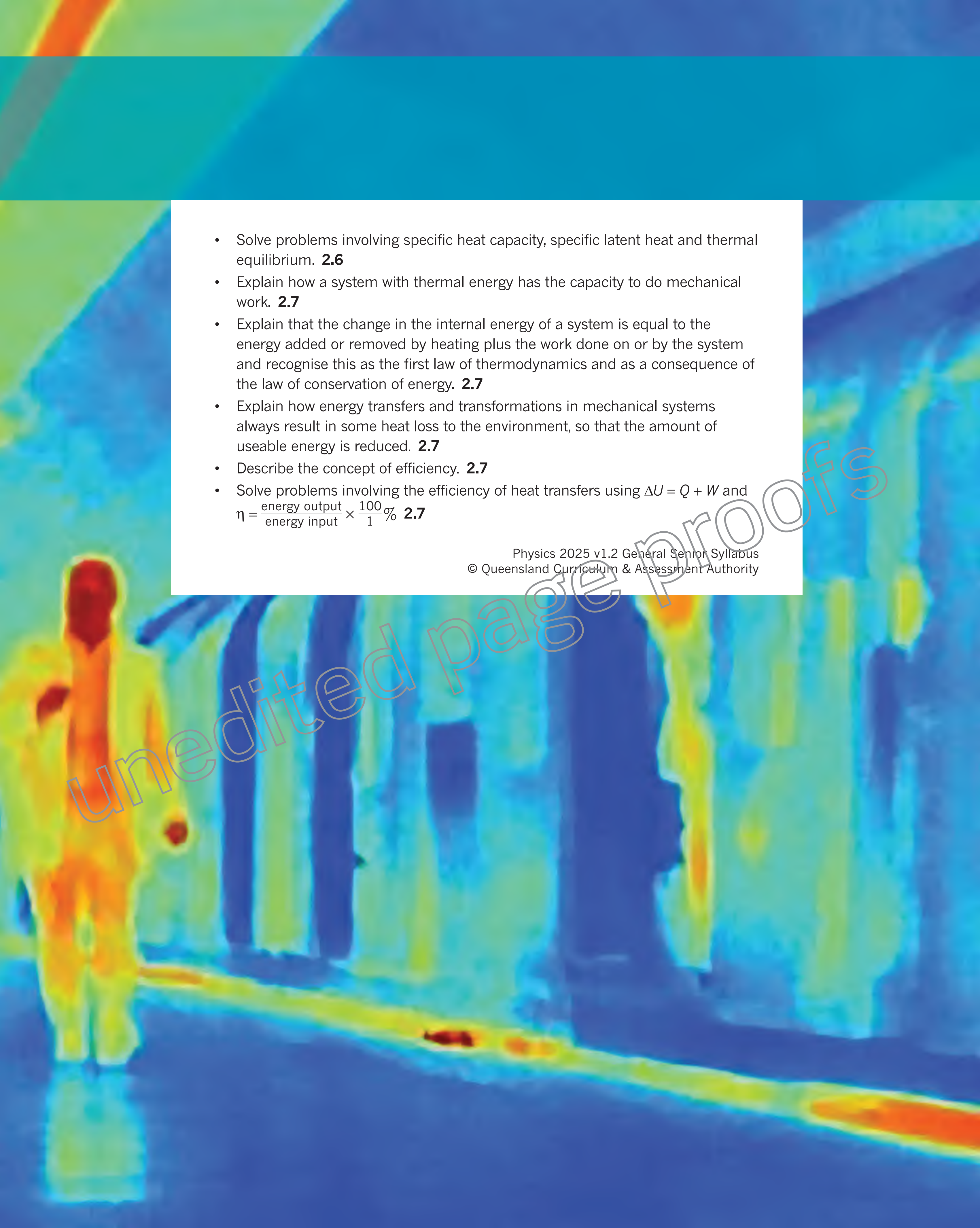
Topic 1 • Heating processes

KINETIC PARTICLE MODEL AND SPECIFIC HEAT CAPACITY

- Describe the kinetic particle model of matter. **2.1**
- Describe the concepts of thermal energy, temperature, kinetic energy, heat and internal energy. **2.1**
- Explain heat transfers in terms of conduction, convection and radiation. **2.3**
- Use $T_K = T_C + 273$ to convert temperature measurements. **2.2**
- Explain that a change in temperature is due to the addition or removal of energy from a system (without phase change). **2.2**
- Describe the concept of specific heat capacity. **2.4**
- Solve problems involving specific heat capacity using $Q = mc\Delta T$. **2.4**
- Interpret data from specific heat capacity experiments. **2.4**

PHASE CHANGES AND ENERGY CONSERVATION

- Explain, in terms of the internal energy of a system and the kinetic particle model of matter, why the temperature of a system remains the same during the process of state change. **2.5**
- Describe the concept of specific latent heat. **2.5**
- Solve problems involving specific latent heat using $Q = mL$. **2.5**
- Describe the concept of thermal equilibrium in terms of the temperature and average kinetic energy of the particles in each of the systems. **2.6**
- Explain the process in which thermal energy is transferred between two systems until thermal equilibrium is achieved and recognise this as the zeroth law of thermodynamics. **2.6**



- Solve problems involving specific heat capacity, specific latent heat and thermal equilibrium. **2.6**
- Explain how a system with thermal energy has the capacity to do mechanical work. **2.7**
- Explain that the change in the internal energy of a system is equal to the energy added or removed by heating plus the work done on or by the system and recognise this as the first law of thermodynamics and as a consequence of the law of conservation of energy. **2.7**
- Explain how energy transfers and transformations in mechanical systems always result in some heat loss to the environment, so that the amount of useable energy is reduced. **2.7**
- Describe the concept of efficiency. **2.7**
- Solve problems involving the efficiency of heat transfers using $\Delta U = Q + W$ and $\eta = \frac{\text{energy output}}{\text{energy input}} \times \frac{100}{1}\%$ **2.7**

2.1 The kinetic particle model

BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- describe the kinetic particle model of matter
- apply the kinetic particle model to solids, liquids and gases
- understand potential and kinetic energy
- understand heat, temperature and thermal energy
- recognise the internal energy of a substance is a combination of its potential and kinetic energy.

In the 16th century, Sir Francis Bacon, an English essayist and philosopher, proposed a radical idea: that heat is motion. He went on to write that heat is the rapid vibration of tiny particles within every substance. At the time, his ideas were dismissed because the nature of particles wasn't fully understood. An opposing theory at the time was that heat was related to the movement of a fluid called 'caloric' that filled the spaces within a substance.

Today, it is understood that all matter is made up of small particles (atoms or molecules). Using this knowledge, it is possible to look more closely at what happens during heating processes. This section starts by looking at the kinetic particle model, which describes the behaviour of the particles in a substance.

KINETIC PARTICLE MODEL OF MATTER

A **model** is a representation that describes or explains the workings within an object, system or idea. Models can be used to simplify complex structures or systems. Scientists use models to understand and analyse complex systems, and to make predictions about these systems.

The behaviour of particles in a substance is complex, so a model is used to describe it.

Some philosophers in the 17th century believed that heat was a fluid that filled the spaces between the particles of a substance and flowed from one substance to another. This is known as the 'caloric' theory. When caloric flowed from one substance into another, the first object cooled down and the second object heated up. Many attempts were made to detect caloric, but none were successful. It was assumed that caloric had no mass, odour, taste or colour. Scientists now know that caloric simply doesn't exist.

Much of our understanding of the behaviour of matter today depends on a model called the **kinetic particle model** (also known as kinetic theory).

The kinetic particle model of matter states that the small particles (atoms or molecules) that make up all matter have kinetic energy. This means that all particles are in constant motion, even in extremely cold solids. It was thought centuries ago that if a material was continually made cooler, there would be a point at which the particles would eventually stop moving. This coldest possible **temperature** is called **absolute zero** and will be discussed later in Module 2.2.

These are the assumptions of the kinetic particle model:

- All matter is made up of many very small particles (atoms or molecules).
- The particles are in constant motion.
- No kinetic energy is lost or gained overall during collisions between particles.
- There are forces of attraction and repulsion between the particles in a material.
- The distances between particles in a gas are large compared with the size of the particles.

The kinetic theory applies to all states (or phases) of matter: solids, liquids, gases and plasmas.

Solids

In a solid, the matter holds its fixed shape because the particles exert forces on each other. There are also repulsive forces, without which the attractive forces would cause the solid to collapse. In a solid, the attractive and repulsive forces hold these particles in more or less fixed positions, usually in a regular arrangement or lattice, as shown in Figure 2.1.1a. The particles in a solid are not completely still; they vibrate around average positions. The forces on individual particles are sometimes predominantly attractive and sometimes repulsive, depending on their exact position relative to neighbouring particles.

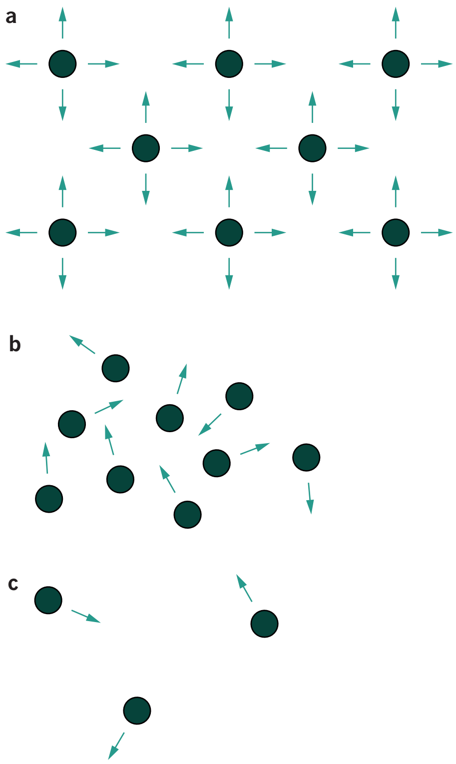


FIGURE 2.1.1 (a) Molecules in a solid have low kinetic energy and vibrate around average positions within a regular arrangement. (b) The molecules in a liquid have more kinetic energy than those in a solid. They move more freely and will take the shape of the container. (c) Gas molecules are free to move in any direction.

Liquids

In a liquid, the particles exert a balance of forces of attraction and repulsion. Compared with a solid, the particles in a liquid have more freedom to move around each other and will therefore take the shape of the container. Particles collide but remain attracted to each other, so the liquid remains within a fixed volume but with no fixed shape (Figure 2.1.1b). In general, a liquid takes up a slightly greater volume than the same amount of matter would in the solid state. One exception is water, which is less dense as a solid (ice).

Gases

In a gas, particles are in constant, random motion, colliding with each other and the walls of the container. The particles move rapidly in every direction, quickly filling the volume of any container and occasionally colliding with each other (Figure 2.1.1c). A gas has no fixed volume. The particle speeds are high enough that, when the particles collide, the attractive forces are not strong enough to keep the particles close together. The repulsive forces cause the particles to separate and move off in other directions.

Plasmas

Plasma exists when matter is heated to very high temperatures and electrons are freed (ionisation). A gas that is ionised and has an equal number of positive and negative charges is called plasma. The interior of stars consists of plasma. In fact, most of the matter in the universe is plasma (Figure 2.1.2).

FIGURE 2.1.2 99.9% of the visible universe is made up of plasma.

WS 1.1.2

ENERGY

Energy is a very important concept in the study of the physical world,and is a focus in all areas of scientific study. Later chapters investigate energy in more detail.

Work is defined as being done when a force is applied to an object and moves it.

Energy is a measure of an object's ability to do work. Work is done when energy is transferred or transformed. For example, raising an object's temperature or moving or lifting an object is referred to as doing work. Work is measured in joules. The symbol for joules is J. Figure 2.1.3 shows the amount of energy available from some energy sources.

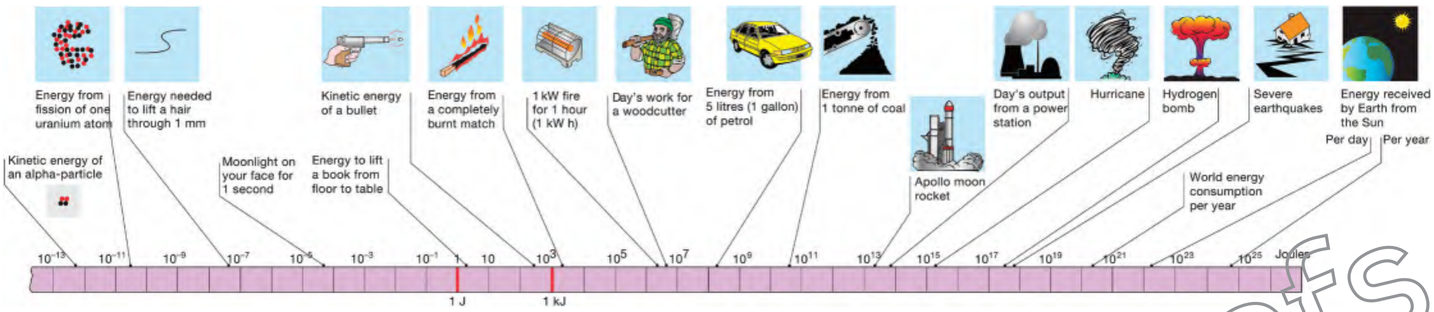


FIGURE 2.1.3 The comparative amounts of energy available from several sources.

Potential energy

Potential energy is stored energy. Examples of potential energy include gravitational, nuclear, spring and chemical energy. Chemical potential energy is associated with the bonds between the particles of a substance. An increase in the potential energy of particles in a substance results in movement of the particles from their equilibrium positions.

Kinetic energy

Kinetic energy is the energy of movement. It is equal to the amount of work needed to bring an object from rest to its present speed or to return it to rest.

Internal energy and heat

The kinetic particle model can be used to explain the idea of heat as a transfer of energy. **Heat** (measured in joules) is the transfer of **thermal energy** from a hotter to a colder body. Thermal energy is the energy a substance has because of its temperature. Heating is observed by the change in temperature, the change of state or the expansion of a substance. When a substance is 'heated', the particles within the material gain kinetic energy and/or potential energy. Conversely, when the average kinetic energy of the particles in a substance increases, the temperature of the substance also increases.

The term 'heat' refers to energy that is being transferred (moved), so it is incorrect to talk about heat contained in a substance. The term **internal energy** refers to the total kinetic and potential energy of the particles within a substance. Heating (the transfer of thermal energy) changes the internal energy of a substance by affecting the kinetic energy and/or potential energy of the particles within the substance. Individual particles in a substance move back and forth in an ordered manner, due to their kinetic energy, and this behaviour can be modelled. In contrast, the internal energy of a system is associated with the chaotic motion of the particles and relates to the behaviour of many particles that all have their own kinetic and potential energy.

- Heating is a process that always transfers thermal energy from a hotter substance to a colder substance.
- Heat is measured in joules (J).
- Temperature is related to the average kinetic energy of the particles in the substance. The faster the particles move, the higher the temperature of the substance.

According to the kinetic particle model, an increase in the total internal energy of the particles in a substance will result in an increase in temperature if there is a net gain in kinetic energy. Hot air balloons are an example of this process in action. The air in a hot air balloon is heated by a gas burner to a maximum of 120°C. The nitrogen and oxygen molecules in the air in the balloon gain energy and so move faster. The air in the balloon becomes less dense than the surrounding air, causing the balloon to float, as shown in Figure 2.1.4.

FIGURE 2.1.4 (a) Nitrogen and oxygen molecules gain energy when the air is heated, lowering the density of the air and causing the hot air balloon to rise off the ground. (b) A thermal image shows the temperature of the air inside the balloon.

Sometimes heating results only in the change of state or expansion of an object (phase change), and not a change in temperature. In these cases, the total internal energy of the particles has increased, but only the potential energy has increased; kinetic energy has not changed.

For example, particles in a solid being heated will continue to be mostly held in place, due to the relatively strong interparticle forces. For the substance to change state from solid to liquid, it must receive enough energy to separate the particles from each other and disrupt the regular arrangement of the solid. During this 'phase change' process, the energy is used to overcome the strong interparticle forces but does not change the overall speed of the particles. In this situation, the temperature does not change. This will be discussed in more detail in Module 2.5.

Kinetic energy and temperature

As the temperature of a substance increases, the kinetic energy and therefore the speed of the particles also increases. At a particular temperature, the particles in a substance will have a range of kinetic energies. Although most of the particles have similar energies, there are always some particles with a high energy or a low energy. This range of energies is shown on a graph called a **Maxwell–Boltzmann distribution**. Figure 2.1.5 shows how the range of energies is represented in a Maxwell–Boltzmann distribution.

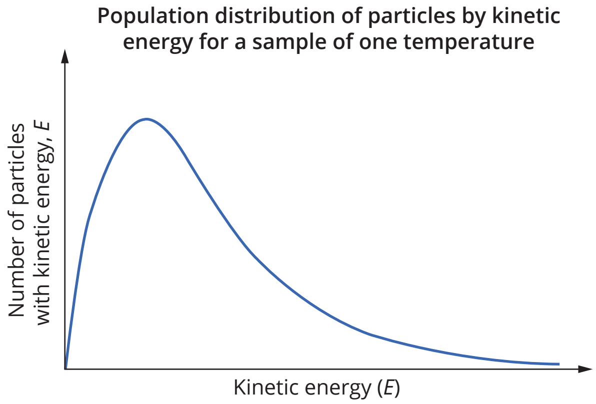


FIGURE 2.1.5 The Maxwell–Boltzmann distribution shows the range of kinetic energies of particles in a sample at a particular temperature. The peak of the graph corresponds to the energy of the greatest number of particles.

As the temperature of a system increases, the average kinetic energy of the particles increases and therefore the average velocity of the particles also increases. Figure 2.1.6 shows the distribution of kinetic energies for a gas at three different temperatures. Notice that the area under the curve, which is equal to the total number of particles in the sample, stays constant when the temperature is changed. As the temperature increases, the increasing average kinetic energy of the particles can be seen by the movement to the right of the peak in the Maxwell–Boltzmann distribution.

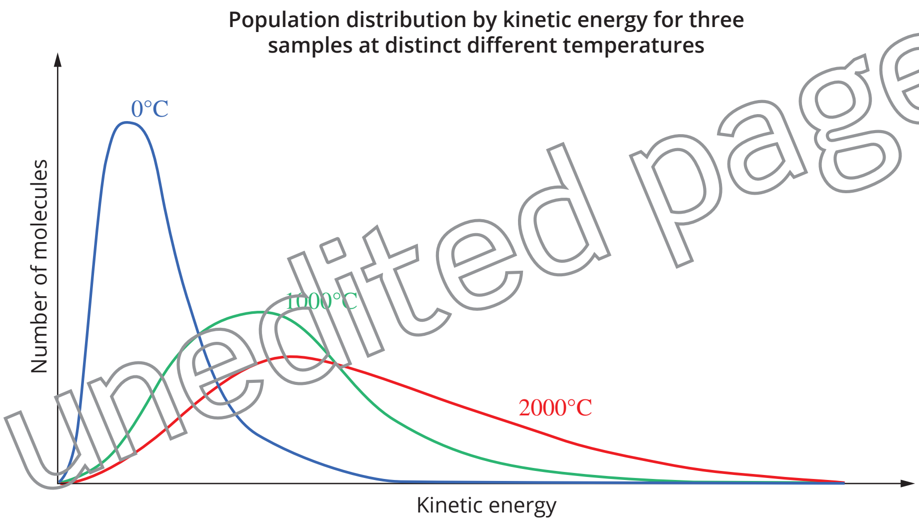


FIGURE 2.1.6 The Maxwell–Boltzmann kinetic energy distribution for a range of temperatures.

2.1 Review

SUMMARY

- The kinetic particle theory proposes that all matter is made of atoms or molecules (particles) that are in constant motion.
- In solids, attractive and repulsive forces hold the particles in more-or-less fixed positions, usually in a regular arrangement or lattice. These particles are not completely still; they vibrate about average positions.
- In liquids, there is still a balance of attractive and repulsive forces between particles, but the particles have more freedom to move around. Liquids maintain a fixed volume.
- In gases, the particle speeds are high enough that, when particles collide, the attractive forces are not strong enough to keep them close together. The repulsive forces cause the particles to move off in other directions.
- Internal energy refers to the total kinetic and potential energy of the particles within a substance.
- Temperature is related to the average kinetic energy of the particles in a substance.
- Heating is a process that always transfers thermal energy from a hotter substance to a colder substance.

KEY QUESTIONS

Describe

1 What are the assumptions of the kinetic particle model?

2 Describe the two forms of internal energy.

3 What does the Maxwell–Boltzmann distribution show?

Apply

4 Explain how you measure the change of the average kinetic energy of the molecules within a substance.

5 Describe which way the heat will flow when you put some hot coffee into a cold mug.

6 Explain how a hot air balloon can rise.

7 Use the kinetic particle model to explain why it is hard to compress a solid.

8 Explain how liquids can change shape using the kinetic particle model.

9 Why is it incorrect to talk about how much heat an object has?

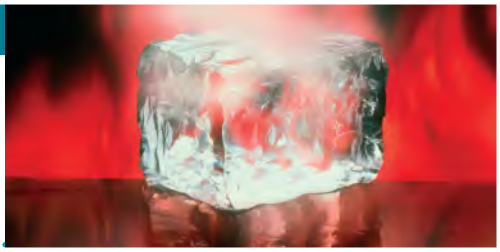
Analyse

10 Use the kinetic theory of matter to differentiate between water ice, liquid water and water vapour.

2.2 Temperature

BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- understand the difference between absolute and arbitrary scales
- convert temperature measurements between Celsius and Kelvin
- explain how a change in temperature is due to an addition or removal of energy from a system.



MEASURING TEMPERATURE

Before thermometers were invented, temperatures were described using vague terms such as hot, cold and lukewarm. In about 1593, Italian inventor Galileo Galilei made one of the first thermometers. His 'thermoscope' was not particularly accurate as it did not consider changes in air pressure, but it did suggest some basic principles for determining a suitable scale of measurement. His work suggested that there be two fixed points: the hottest day of summer and the coldest day of winter. A scale like this is referred to as an arbitrary scale, because the fixed points are randomly chosen.

How to measure temperature

To measure temperature, you use properties of materials that change with temperature. Many thermometers, including mercury and alcohol thermometers (Figure 2.2.1), use thermal expansion. Thermal expansion is the expansion of a substance as it increases in temperature. The mercury (or alcohol) expands as it gets warmer, filling more of the thermometer, producing a higher temperature reading.

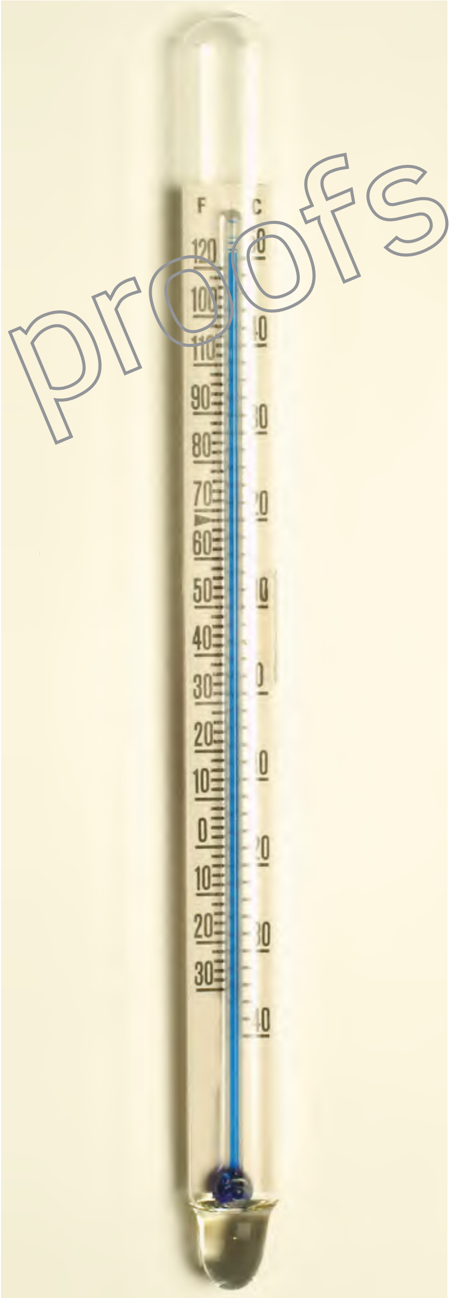


FIGURE 2.2.1 An alcohol thermometer contains alcohol (red or blue, as shown here) instead of mercury (silver).

TEMPERATURE SCALES

Two of the better-known arbitrary temperature scales are Fahrenheit and Celsius. German physicist Gabriel Fahrenheit invented the first mercury thermometer in 1714. The Fahrenheit scale is often used in the USA to measure temperature, but the Celsius scale is used in most other countries of the world.

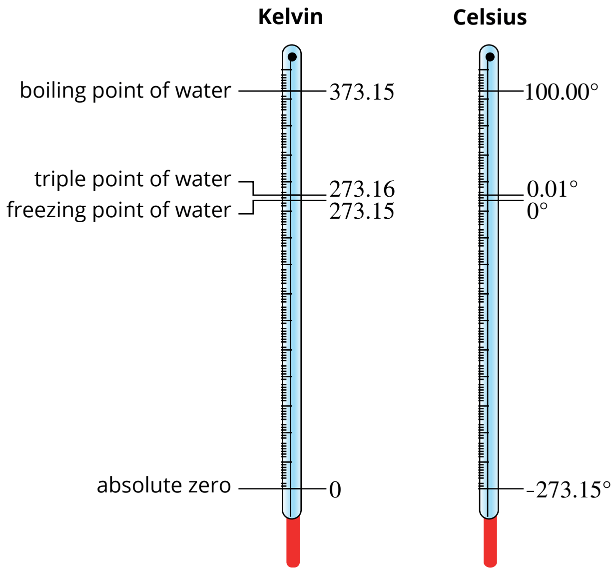
Absolute scales are different from arbitrary scales. For a scale to be regarded as 'absolute', it should have no negative values. The fixed points must be reproducible and have zero as the lowest value.

Kelvin temperature scale

When developing the absolute temperature scale, the triple point of water provides one reliable fixed point. The triple point is the point at which the combination of temperature and air pressure allows all three states of water to coexist. The triple point of water is only slightly above the standard freezing point (0.01°C) and provides a unique and repeatable temperature with which to adjust the Celsius scale.

The absolute or **Kelvin** temperature scale is based on absolute zero and the triple point of water. Figure 2.2.2 shows a comparison of the Kelvin and Celsius scales.

- The freezing point of water (0°C) is equivalent to 273.15 K (kelvin). This is approximately 273 K.
- The size of each unit (1°C or 1 K) is the same.
- The word 'degree' and the degree symbol are not used with the Kelvin scale.
- 0°C is the freezing point of water at standard atmospheric pressure.
- 100°C is the boiling point of water at standard atmospheric pressure.



WS 1.1.3

FIGURE 2.2.2 Comparison of the Kelvin and Celsius scales. Note that there are no negative values on the Kelvin scale.

SKILLBUILDER

Converting temperature units between kelvin and degrees Celsius

The standard (SI) unit for measuring temperature is the kelvin. However, in our day-to-day lives, we commonly use the Celsius scale. So, it is important to be able to convert values from one unit to the other.

To convert a temperature in degrees Celsius, T_C, to a temperature in kelvin, T_K, add 273:

$T_K = T_C + 273$

Converting from degrees Celsius to kelvin

The boiling point of ethanol is 78°C

To convert a temperature from degrees Celsius to kelvin, add 273.

$T_K = T_C + 273$

78°C = 78 + 273 K
= 351 K

Converting from kelvin to degrees Celsius

Polyethylene, used to make plastic bottles and bags, has a melting point of approximately 393 K

To convert a temperature from kelvin to degrees Celsius, subtract 273.

$T_C = T_K - 273$

393 K = 393 − 273°C
= 120°C

Worked example 2.2.1

CONVERTING TEMPERATURE UNITS

(a) Convert 450°C into K	
Thinking	**Working**
State the relationship between degrees Celsius and kelvin.	$T_K = T_C + 273$
Substitute in the values to convert from °C into K.	$T_K = 450 + 273$ $= 723$ K
(b) Convert 189 K into °C	
Thinking	**Working**
State the relationship between degrees Celsius and kelvin.	$T_C = T_K - 273$
Substitute in the values to convert from K into °C.	$T_C = 189 - 273$ $= -84$°C

➤ Try yourself 2.2.1

CONVERTING TEMPERATURE UNITS

Convert the following temperatures.
(a) 20°C into K
(b) 654 K into °C

Absolute zero

Experiments indicate that there is a limit to how cold things can get. The kinetic theory suggests that when a given quantity of gas is cooled, its volume decreases. The volume can be plotted against temperature and results in a straight-line graph, as shown in Figure 2.2.3. These experiments were conducted for many different gases, and they all result in the same temperature when the graph is extrapolated to zero volume. For every ideal gas, zero volume is at –273.15°C. At 0 K, therefore, the internal motion of the particles in an ideal gas is zero, as is the volume.

Changing temperature

The temperature of a substance is a measure of the average kinetic energy of the molecules it contains. To increase the temperature of an object, you need to increase the average kinetic energy of its molecules. When an object is heated, extra energy is added to the molecules of the object, increasing their kinetic energy. The increase in kinetic energy when heating a substance applies only when a phase change is not taking place (see Module 2.5).

To cool an object, the average kinetic energy needs to decrease. One way to remove kinetic energy from an object is to put it in contact with a colder object. The heat will flow from the hotter object to the colder object, cooling the hotter object as it shares its thermal energy. The flow of heat will be explored in Module 2.3.

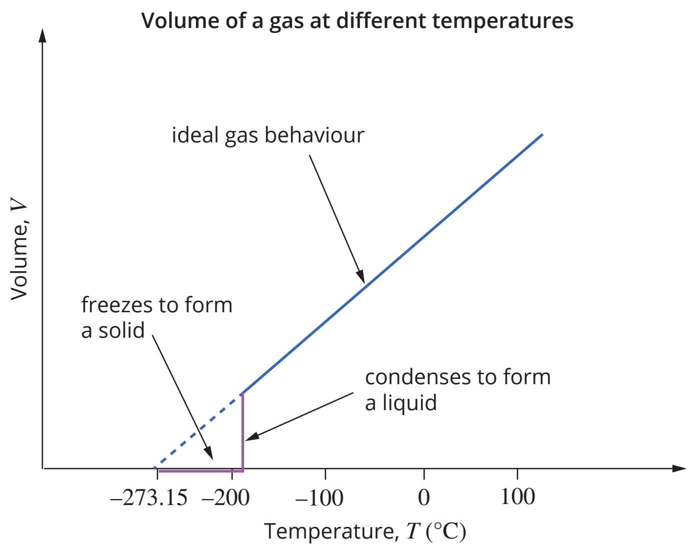


FIGURE 2.2.3 The volume of a gas decreases as the gas cools. This relationship is linear. Extrapolating (extending) the line for any substance to where the volume is zero gives a theoretical value of absolute zero.

2.2 Review

SUMMARY

- Temperatures can be measured in degrees Celsius (°C) or kelvin (K).
- Absolute zero is called simply ‘zero kelvin’ (0 K) and it is equal to –273.15°C.
- The size of each unit, 1°C or 1 K, is the same.
- To convert from degrees Celsius to kelvin: add 273; to convert from kelvin to degrees Celsius: subtract 273.
- A change in temperature is due to the addition or removal of heat from a substance when a phase change is not taking place.

KEY QUESTIONS

Describe

1 Describe how a mercury thermometer measures temperature.

2 Explain the concept of absolute zero temperature and the Kelvin scale.

Apply

3 Covert the following temperatures:

a 30.0°C into K
b 375 K into °C
c 78°C into K
d 40 K into °C
e 1200°C into K
f 6000 K into °C
g 437°C into K

4 A tank of pure helium is cooled to its freezing point of –272.2°C. Describe the energy of the helium particles at this temperature.

Analyse

5 Sort the following temperatures from coldest to hottest:

freezing point of water
100.0 K
absolute zero
–180°C
10.0 K

6 Tank A is filled with hydrogen gas at 0.0°C and another tank, B, is filled with hydrogen gas at 300.0 K. Contrast the average kinetic energy of the hydrogen particles in each tank.

7 What is the temperature of an object that has twice as much thermal energy as an object that is 15°C.

8 Explain how absolute zero temperature was determined. You may need some internet research to help with this.

2.3 Heat flow

BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- understand conduction, convection and radiation
- explain heat transfers in terms of conduction, convection and radiation

If two objects at different temperatures are in thermal contact (that is, they can exchange energy via heat processes), then thermal energy will transfer from the hotter object to the cooler object. Figure 2.3.1 shows how, by preventing the chick's thermal contact with the cold ice, this adult can protect the vulnerable penguin offspring.

There are three possible means by which heat can be transferred:

- conduction
- convection
- radiation.

FIGURE 2.3.1 Emperor penguin chicks avoid heat loss through conduction by sitting on the adult's feet. In this way they avoid contact with the ice.

CONDUCTION

Conduction is the process by which heat is transferred without the net movement of particles (atoms or molecules) from one area to another. Particles in one object vibrate less, particles in the other object vibrate more, but no particles move from one object to the other. Conduction can occur within a material or between materials that are in thermal contact. For example, if one end of a steel rod is placed in a fire, heat will travel along the rod so that the far end of the rod will also heat up. If a person holds an ice cube, then heat will travel from their hand to the ice.

All materials will conduct heat to some extent, but this process is most significant in solids. Conduction is important in liquids and plays a lesser role in the movement of energy in gases.

Conductors and insulators

Materials that conduct heat readily are referred to as good thermal **conductors**. Materials that are poor thermal conductors are referred to as **insulators**. An example of a good conductor and a good insulator can be seen in Figure 2.3.2.

In secondary physics, the terms 'conductor' and 'insulator' are used in the context of both electricity and heating processes. What makes a material a good conductor of heat doesn't necessarily make it a good conductor of electricity. The two types of conduction are related but it's important not to confuse the two processes. A material's ability to conduct heat depends on how conduction occurs within the material.

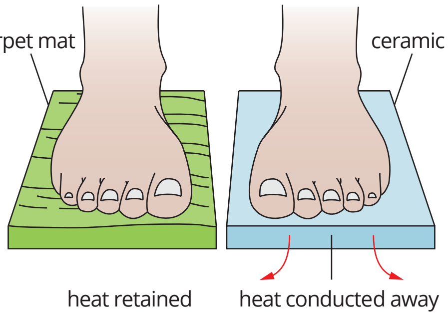


FIGURE 2.3.2 Ceramic floor tiles are good conductors of heat because they conduct heat away from the foot, so the tiles feel cold to your feet. The carpet mat is a thermal insulator because thermal energy from the foot is not transferred away as quickly, so the carpet doesn't feel as cold to your feet.

Conduction can happen in two ways:

- by energy transfer through molecular or atomic collisions
- by energy transfer by free electrons.

Thermal transfer by collision

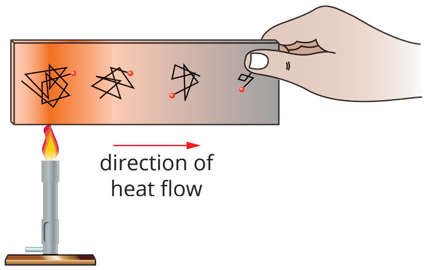


FIGURE 2.3.3 Thermal energy is passed on by collisions between adjacent particles.

The kinetic particle model explains that particles in a solid substance are constantly vibrating within the material structure and so interact with neighbouring particles. If one part of the material is heated, then the particles in that region will vibrate more rapidly. Interactions with neighbouring particles will pass on this kinetic energy throughout the system via the bonds between the particles (Figure 2.3.3).

Heat transfer by collision is a slow process because the mass of the particles is relatively large and the vibrational velocities are fairly low. This method of conduction is used for heat transfer in poor conductors of heat (thermal insulators) such as glass, wood and paper.

Thermal transfer by free electrons

Some materials, particularly metals, have electrons that are not directly involved in any one specific chemical bond. Therefore, these electrons are free to move throughout the lattice of positive ions. These free electrons can carry heat through the material.

For example, if a metal is heated, then not only will the positive ions within the metal gain extra energy but so will these free electrons. As the electron's mass is considerably less than that of the positive ions, even a small energy gain will result in a very large gain in velocity. Consequently, these free electrons provide a means by which heat can be quickly transferred throughout the whole of the material. It is therefore no surprise that metals, which are good electrical conductors because of these free electrons, are also good thermal conductors.

Thermal conductivity

Thermal conductivity describes the ability of a material to conduct heat. It is temperature dependent and is measured in watts per metre kelvin ($W\,m^{-1}\,K^{-1}$). Table 2.3.1 highlights the difference in conductivity between metals and other substances.

TABLE 2.3.1 Thermal conductivities of some common materials

Material	Conductivity ($W\,m^{-1}\,K^{-1}$)
silver	420
copper	380
aluminium	240
steel	60
ice	2.2
brick, glass	≈ 1
concrete	≈ 1 (depending on composition)
water	0.6
human tissue	0.2
wood	0.15
polystyrene	0.08
paper	0.06
fibreglass	0.04
air	0.025

Factors that affect thermal conductivity

The rate at which heat is transferred is measured in joules per second ($J\,s^{-1}$), or watts (W). The rate at which heat will be transferred through a system is dependent on:

- the nature of the material. The larger a material's thermal conductivity, the more rapidly it will conduct heat energy.
- the temperature difference. A greater temperature difference between the two objects will result in a faster rate of energy transfer.

- the thickness of the material. Thicker materials require a greater number of collisions between particles or movement of electrons to transfer energy from one side to the other.
- the surface area. Increasing the surface area relative to the volume of a system increases the number of particles involved in the transfer process, increasing the rate of conduction.

CONVECTION

Convection is the transfer of thermal energy within a fluid (liquid or a gas) by the movement of hot areas from one place to another. Unlike other forms of heat transfer, such as conduction and radiation, convection involves the mass movement of particles within a system over a considerable distance.

Heating by convection

Although liquids and gases are generally not good conductors of thermal energy, heat can be transferred quite quickly through liquids and gases by convection. Unlike other forms of thermal energy transfer, convection involves the mass movement of particles within a system over a distance.

As a fluid is heated, its particles gain kinetic energy and push apart due to the increased vibration of the particles. This causes the density of the heated fluid to decrease and the heated fluid rises. Colder fluid, with slower moving particles, is more dense and heavier and therefore falls, moving in to take the place of the warmer fluid. A convection current forms when there is warm fluid rising and cool fluid falling. This action can be seen in Figure 2.3.4. Upwellings in oceans, wind and weather patterns are at least partially due to convection on a very large scale.

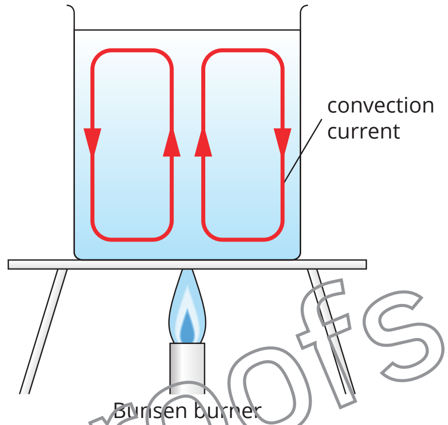


FIGURE 2.3.4 When a liquid or gas is heated, it becomes hotter and less dense so will rise. The colder, denser fluid will fall. As this fluid heats up, it in turn will rise, creating a convection current.

Factors that affect thermal convection

It is difficult to quantify the thermal energy transferred via convection, but some estimates can be made. The rate at which convection will occur is affected by:

- the temperature difference between the heat source and the convective fluid
- the surface area exposed to the convective fluid.

In a container, the effectiveness of convection to transfer heat depends on the placement of the source of heat. For example, the heating element in a kettle is always found near the bottom of the kettle. From this position, convection currents form throughout the water to heat it more effectively (Figure 2.3.5a). If the heating element is placed near the top of the kettle, convection currents form only near the top. This is because the hotter water is less dense than the cooler water below and will remain near the top. Convection currents will not form throughout the water (Figure 2.3.5b).

To visualise convection currents, set up a glass beaker with one side positioned over a Bunsen burner. Fill the beaker with water and add a couple of drops of food colouring. The convection patterns in the water will be highlighted and make a great photograph, as shown in the Figure 2.3.6.

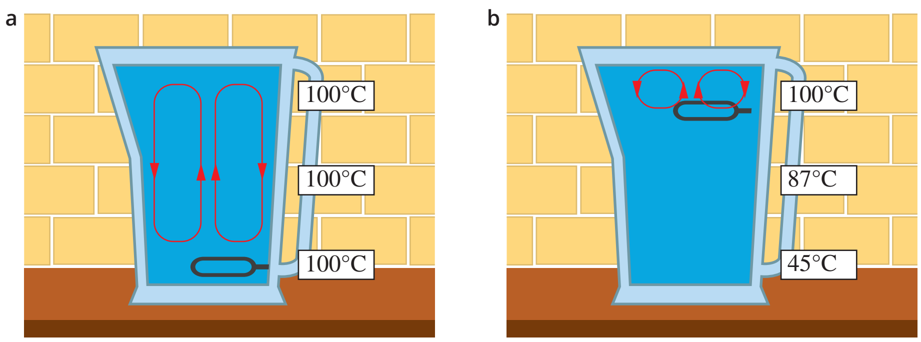


FIGURE 2.3.5 (a) By placing the heating element at the bottom of a kettle, the water near the bottom is heated and rises, forming convection currents throughout the entire depth of the water. (b) If the heating element is placed near the top of the kettle, the convection currents form near the top and heat transfer is slower.

PA
1.1.2

FIGURE 2.3.6 Demonstration of convection using a Bunsen burner, water and food colouring

Applications of convection

There are two main causes of convection:

- forced convection, in which heat transfer is caused by an external source. An example is ducted heating in which air is heated and then blown into a room. Applications of forced convention are considered when designing heating and cooling systems.
- natural convection, in which heat transfer is induced by different temperature gradients of fluids in nature. Ocean currents are an example of natural convection.

A dramatic example of natural convection is the thunderhead clouds of summer storms (Figure 2.3.7), which form when hot, humid air from natural convection currents is carried rapidly upwards into the cooler upper atmosphere.

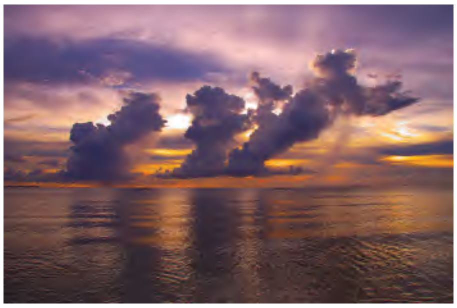


FIGURE 2.3.7 The thunderheads of summer storms are a very visible indication of natural convection in action.

Paragliders fly by sitting in a harness suspended beneath a fabric wing (Figure 2.3.8). They gain altitude by catching thermals. Thermals are columns of rising hot air created by dark regions on the ground that have been heated up by the Sun. Roads, rock faces and ploughed fields are good at creating thermals.

FIGURE 2.3.8 Paragliders can gain altitude by finding a thermal. Thermals are areas of rising hot air created by hot regions on the ground.

RADIATION

Both convection and conduction involve the transfer of heat through matter. Life on Earth depends upon the transfer of energy from the Sun through the near-vacuum of space. If heat could only be transferred by the action of particles, then the Sun's energy would never reach Earth. Radiation is a means of transfer of heat without the movement of matter.

Electromagnetic radiation

In this context, **radiation** is a shortened form of electromagnetic radiation, which includes visible, ultraviolet and infrared light. Together with other forms of light, these make up the **electromagnetic spectrum**.

Heat is transferred from one place to another without the movement of particles by electromagnetic radiation (light). Electromagnetic radiation travels at the speed of light. When electromagnetic radiation hits an object, it will be partially reflected, partially transmitted and partially absorbed. The absorbed radiation transfers thermal energy to the absorbing object and causes a rise in temperature. When you hold a marshmallow by an open fire, you are using a combination of radiation and convection to toast the marshmallow (Figure 2.3.9).

FIGURE 2.3.9 Heat transfer from the flame to the marshmallow is an example of radiation.

Electromagnetic radiation is emitted by all objects that are at a temperature above absolute zero (0 K or −273°C). The **wavelength** and **frequency** of the emitted radiation depend on the internal energy of the object. The higher the temperature of the object, the higher the frequency and the shorter the wavelength of the radiation emitted. The total radiant energy emitted increases as the temperature of the system increases. This can be seen in Figure 2.3.10.

A human body emits radiation in the infrared range of wavelengths, whereas hotter objects emit radiation of a higher frequency and shorter wavelength. Hotter objects can emit radiation in the range of visible, ultraviolet and shorter wavelengths of the electromagnetic spectrum. For example, as a red-hot fire poker heats up further, it becomes yellow-hot. Wavelengths that you feel as warm tend to be below 9.5μm.

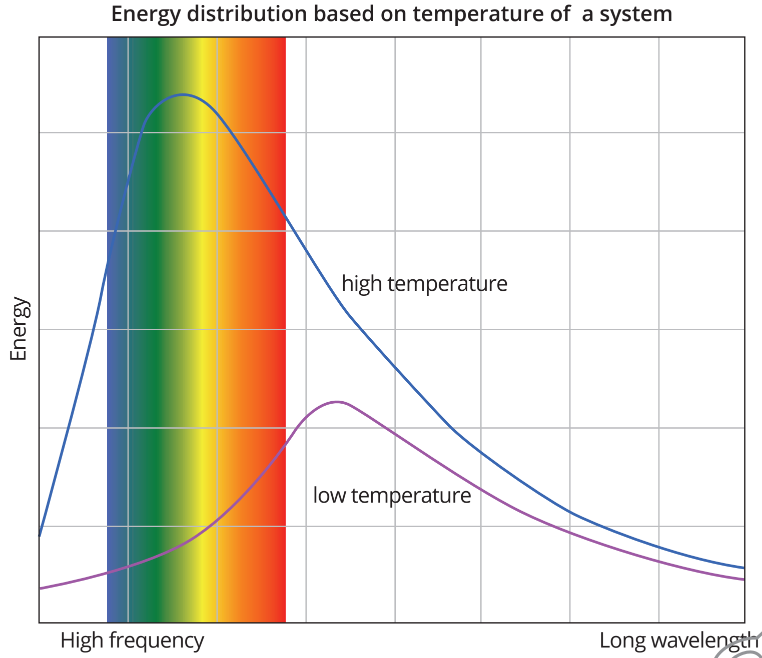


FIGURE 2.3.10 A system emits radiation over a range of frequencies. At a low temperature, it will emit small amounts of radiation of longer wavelengths. As the temperature of the system increases, more short-wavelength radiation is emitted and the total radiant energy emitted increases.

Absorption and emission of radiant energy

All objects both absorb and emit thermal energy by radiation. If an object absorbs more thermal energy than it emits, its temperature will increase. If an object emits more energy than it absorbs, its temperature will decrease. If no temperature change occurs, the object and its surroundings are in thermal equilibrium.

All objects emit some radiation, but they will not all emit or absorb at the same rate. Several factors affect both the rate of **emission** and the rate of **absorption**.

- Surface area: The larger the exposed surface area, the higher the rate of radiant transfer.
- Temperature: The greater the difference between the temperature of the absorbing or emitting surface and the temperature of its surroundings, the greater the rate of energy transfer by radiation.
- Wavelength of the **incident** radiation: Matte black surfaces are almost perfect absorbers of radiant energy at all wavelengths. Highly reflective surfaces are good reflectors of all wavelengths. An example of how reflective surfaces can be exploited is shown in Figure 2.3.11. For all other surfaces, the absorption of particular wavelengths of radiant energy will be affected by the wavelength of that energy. For example, white surfaces absorb visible wavelengths of radiant energy poorly, but they will absorb infrared radiation just as well as black surfaces do.
- Surface colour and texture: The characteristics of the surface itself determine how readily that surface will emit or absorb radiant energy. Matte black surfaces will absorb and emit radiant energy faster than shiny, white surfaces. This means that a roughened, dark surface will heat up faster than a shiny, light surface. Matte black objects will also cool down faster because they will radiate energy just as efficiently as they absorb it. Car radiators are painted black to increase the emission by radiation of the thermal energy that is collected from the car engine.

FIGURE 2.3.11 The silvered surface of an emergency blanket reflects thermal energy back to the body, and retains the radiant energy, which would normally be lost. This simple method works as excellent thermal insulation.

2.3 Review

SUMMARY

- Conduction is the process of heat transfer within a material or between materials without the overall transfer of the substance itself.
- All materials will conduct heat to a greater or lesser degree. Materials that readily conduct heat are called good thermal conductors. Materials that conduct heat poorly are called thermal insulators.
- The rate of conduction depends on the temperature difference between two materials, the thickness of the material, the surface area and the nature of the material.
- Convection is the transfer of heat within a fluid (liquid or gas).
- Convection involves the mass movement of particles within a system over a distance.
- A convection current forms when there is warm fluid rising and cool fluid falling.
- Any object whose temperature is greater than absolute zero emits thermal energy by radiation.
- Radiant transfer of thermal energy from one place to another occurs by means of electromagnetic waves.
- When electromagnetic radiation falls on an object, it will be partially reflected, partially transmitted and partially absorbed.
- The rate of emission or absorption of radiant heat will depend on the:
  - temperature difference between the object and the surrounding environment
  - surface area and surface characteristics of the object
  - wavelength of the radiation.

KEY QUESTIONS

Describe

1 List the properties of a material that affect its ability to conduct heat.

2 List the states of matter in which convection can occur.

3 What is the initial direction of the transfer of heat in a convection current?

4 Describe how we feel heat transferred from a fire.

Apply

5 Explain why the process of conduction by molecular collision is slow.

6 Pilots of glider aircraft or hang-gliders, some birds such as eagles and some insects rely on 'thermals' to give them extra lift. Explain how these rising columns of air are established.

7 Explain how a down(feather)-filled quilt keeps a person warm in winter.

8 Why is the vacuum of space a good insulator?

Analyse

9 On a cold day, the plastic or rubber handles of a bicycle feel much warmer than the metal surfaces. Explain this in terms of the thermal conductivity of each material.

10 Convection is referred to as a method of heat transfer through fluids. Evaluate whether it is possible for solids to pass on their heat energy by convection.

11 On a hot day, the top layer of water in a swimming pool can heat up while the lower, deeper parts of the water can remain quite cold. Explain, using the concept of convection, why this happens.

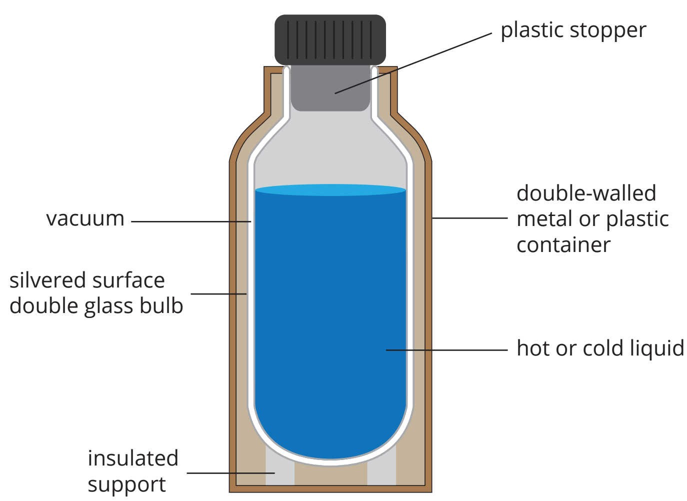
12 Three identical, sealed beakers are filled with near-boiling water. One beaker is painted matte black, one is dull white and the third is gloss white.

   a Predict which beaker will cool fastest.

   b Predict which beaker will cool slowest.

13 A vacuum flask has a tight-fitting stopper at the top. The double glass bulb has silvered inner surfaces and has had most of the air evacuated from it. It is located inside a protective outer case with insulated supports. Assess how this design makes a vacuum flask good at keeping liquids hot inside.

Vacuum or Dewar flask



2.4 Specific heat capacity

BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- understand that the amount of energy required to raise the temperature of a substance depends on the mass of the substance
- understand that the amount of energy required to raise the temperature of a substance depends on what substance it is
- understand that the amount of energy required to raise the temperature of a substance depends on the state of the substance
- analyse, compare and predict the amount of energy required to raise the temperature of a substance
- understand the concept of specific heat capacity.

A small volume of water in a kettle will experience a greater change in temperature than a larger volume, if heated for the same time. A metal object left in the sunshine gets hotter faster than a wooden object. Large heaters warm rooms faster than small heaters.

These simple observations suggest that the mass, material and amount of energy transferred influence any change of temperature.

CHANGING TEMPERATURE OF A SUBSTANCE

The temperature of a substance is a measure of the average kinetic energy of the particles inside the substance. To increase the temperature of the substance, the kinetic energy of its particles must increase. This happens when heat is transferred to that substance. The amount the temperature increases, for a given amount of heat energy, depends on the type of substance and its mass.

The internal structure of a substance determines its **specific heat capacity (*c*)**, which is a measure of how easy it is to heat it up. The specific heat capacity of a material changes when the material changes state. For a given material, therefore, you can calculate how much heat energy is needed to raise the temperature by a given amount using the specific heat capacity, provided a change of state does not occur

Table 2.4.1 lists the specific heat capacities for some common materials. Included in the list is the average value for the human body, which considers the various materials within the body and the proportion that each material contributes to the body's total mass.

Heat energy is calculated from the specific heat capacity for a given temperature change by using the equation:

$Q = mc\Delta T$

where

Q is the heat energy transferred in joules (J)

m is the mass in kilograms (kg)

c is the specific heat capacity of the material ($\mathrm{J\,kg^{-1}\,K^{-1}}$)

Δ means "change in"

ΔT is the change in temperature (°C or K).

TABLE 2.4.1 Approximate specific heat capacities of common substances

Material	c ($\mathrm{J\,kg^{-1}\,K^{-1}}$)
lead	130
mercury	140
brass	370
copper	390
iron	440
glass	840
aluminium	900
air	1000
steam (water)	2000
ice (water)	2100
ethanol	2460
methylated spirits	2500
human body	3500
liquid water	4200

The specific heat capacity of a material, c, is the amount of energy that must be transferred to change the temperature of 1 kg of the material by 1°C or 1 K.

Worked example 2.4.1

CALCULATIONS USING SPECIFIC HEAT CAPACITY

A hot water tank contains 135 L of water. Initially the water is at 20.0°C. Calculate the amount of energy that must be transferred to the water to raise the temperature to 70.0°C.

Thinking	Working
Calculate the mass of water. 1 L of water = 1 kg	Volume = 135 L Mass = 135 kg
ΔT = final temperature – initial temperature	$\Delta T = 70.0 - 20.0$ $= 50.0$°C
From Table 2.4.1: $c_{water} = 4200\ \text{J kg}^{-1}\ \text{K}^{-1}$. Use the equation $Q = mc\Delta T$.	$Q = mc\Delta T$ $= 135 \times 4200 \times 50.0$ $= 28\,350\,000\ \text{J}$ $= 28.4\ \text{MJ}$

➤ Try yourself 2.4.1

CALCULATIONS USING SPECIFIC HEAT CAPACITY

A bath contains 75 L of water. Initially the water is at 50.0°C. Calculate the amount of energy that must be transferred from the water to cool the bath to 30.0°C.

Worked example 2.4.2

COMPARING SPECIFIC HEAT CAPACITIES

Different states of matter of the same substance have different specific heat capacities. Calculate the ratio of the specific heat capacity of liquid water to that of ice.

Thinking	Working
See Table 2.4.1 for the specific heat capacities of water in different states.	$c_{water} = 4200\ \text{J kg}^{-1}\ \text{K}^{-1}$ $c_{ice} = 2100\ \text{J kg}^{-1}\ \text{K}^{-1}$
Divide the specific heat of water by the specific heat of ice. Note that ratios have no units because the units of the two quantities are the same and cancel out.	Ratio $= \frac{c_{water}}{c_{ice}}$ $= \frac{4200}{2100}$ $= 2$

➤ Try yourself 2.4.2

COMPARING SPECIFIC HEAT CAPACITIES

Calculate the ratio of the specific heat capacity of liquid water to that of steam.

SPECIFIC HEAT CAPACITY OF WATER

Note the high specific heat capacity of water in Table 2.4.1. It is about 10 times higher than that of most metals listed. In fact, the specific heat capacity of water is higher than those of most common materials. As a result, water makes a very useful cooling and heat storage agent and is used in areas such as generator cooling towers and car-engine radiators.

Life on Earth also depends on the specific heat capacity of water. About 70% of the Earth's surface is covered by water, and these water bodies can absorb large quantities of thermal energy without great changes in temperature. Oceans both heat up and cool down more slowly than the land areas next to them. This helps to maintain a relatively stable range of temperatures for life on Earth.

Scientists are now monitoring the temperatures of the deep oceans to determine how the ability of oceans to store large amounts of energy might affect climate change. More than 90% of the excess heat trapped in Earth's atmosphere, due to human-caused global warming, is trapped in the oceans.

WS 1.1.4 PA 1.1.1

2.4 Review

SUMMARY

- The specific heat capacity, c, of a substance is a measure of the amount of energy that must be transferred to change the temperature of 1 kg of material by 1°C or 1 K. Its units are $\text{J kg}^{-1}\text{K}^{-1}$.
- When heat is transferred to or from a system or object, the temperature change depends upon the amount of energy transferred, the mass of the material(s) and the specific heat capacity of the material(s): $Q = mc\Delta T$.
- A substance will have different specific heat capacities in different states (solid, liquid, gas).

KEY QUESTIONS

Describe

1 Equal masses of water and aluminium are heated through the same temperature range. Identify the material that requires the most energy to achieve this result. (c_{water} = 4200 J kg^{-1} K^{-1}, $c_{aluminium}$ = 900 J kg^{-1} K^{-1})

2 Determine which has more thermal energy: 10.0 kg of iron at 20.0°C or 10.0 kg of aluminium at 20.0°C. (c_{iron} = 440 J kg^{-1} K^{-1}, $c_{aluminum}$ = 900 J kg^{-1} K^{-1})

Apply

3 100.0 mL of water is heated to change its temperature from 15.0°C to 20.0°C. Calculate how much energy is transferred to the water to achieve this temperature change. (c_{water} = 4200 J kg^{-1} K^{-1})

4 15 400 J of heat energy is transferred to a 12 kg block of iron. Calculate the temperature increase of the iron. (c_{iron} = 440 J kg^{-1} K^{-1})

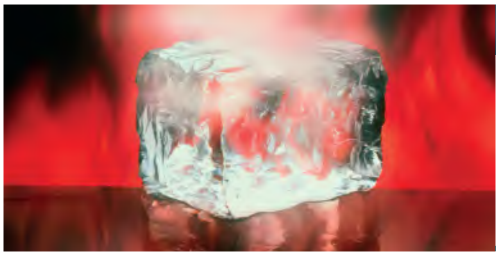
5 A piece of glass is heated with 2688 J of energy. How much glass needs to be heated for a temperature change of 5.00 K? (c_{glass} = 840 J kg^{-1} K^{-1})

Analyse

6 The specific heat capacity of sand is 830 J kg^{-1} K^{-1}. Explain why at the beach on a hot day the sand feels so much hotter than the water. (c_{water} = 4200 J kg^{-1} K^{-1})

7 If 4.0 kJ of energy is required to raise the temperature of 1.0 kg of paraffin by 2.0°C, calculate how much energy (in kJ) is required to raise the temperature of 5.0 kg of paraffin by 1.0°C.

2.5 Phase changes and latent heat



BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- understand changes of state
- understand latent heat
- calculate the energy required to change a substance's state
- interpret a phase change diagram.

If water is heated, its temperature will rise. If enough energy is transferred to the water, eventually the water will boil. The water changes state (from liquid to gas). The **latent heat** is the energy released or absorbed during a change of state. Latent means hidden or unseen. While a substance is in the process of changing state, its temperature remains constant. For example, the energy used when ice melts into water is 'hidden' in the sense that the temperature doesn't rise while the change of state is occurring.

ENERGY AND CHANGE OF STATE

Look at the example of a heating graph for water shown in Figure 2.5.1. In this experiment, heat energy is added at a constant rate, but the increase in the temperature of the water is not always constant. In some sections the temperature increases at a constant rate, but in other sections the temperature remains unchanged (the horizontal sections). In the sections of constant temperature, the material is changing state. The temperature remains constant during the change in state from ice to liquid water and again from liquid water to steam.

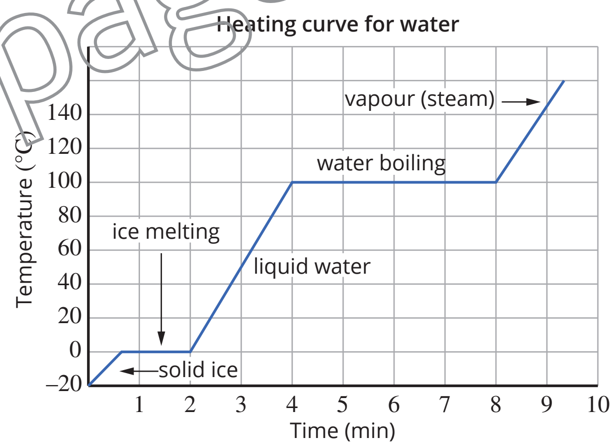


FIGURE 2.5.1 A heating curve for water

LATENT HEAT

The energy needed to change the state of a substance (e.g. solid to liquid, liquid to gas) is called latent heat. Latent heat is the 'hidden' energy that has to be added or removed from a material in order for the material to change state. The energy required, Q, to change the state of a substance of mass m, can be calculated using the latent heat of the substance, L.

$$Q = mL$$

where

Q is the heat energy transferred in joules (J)
m is the mass in kilograms (kg)
L is the latent heat (J kg^{-1}).

The latent heat is calculated using the equation:

$Q = mL$

where

Q is the heat energy transferred in joules (J)
m is the mass in kilograms (kg)
L is the latent heat (J kg^{-1}).

Latent heat of fusion (melting)

The temperature of a solid increases as thermal energy is transferred to the solid. The particles within the solid gain internal energy (as kinetic energy and some potential energy) and their speed of vibration increases. At the point where the solid begins to melt, the particles move further apart, reducing the strength of the bonds holding them in place. At this point, instead of increasing the temperature, the extra energy increases the potential energy of the particles, reducing the interparticle or intermolecular forces. No change in temperature occurs because all the extra energy supplied is used in reducing these forces between particles.

The amount of energy required to melt a solid is the same as the amount of potential energy released when the liquid re-forms into a solid. It is termed the **latent heat of fusion**.

The amount of energy required will depend on the solid. The energy required, Q, to change a solid of mass m to a liquid can be calculated using the latent heat of fusion, L_{fusion}.

$$Q = mL_{fusion}$$

It takes almost 80 times as much energy to turn 1 kg of ice into water (with no temperature change) as it does to raise the temperature of 1 kg of water by 1°C. It takes a lot more energy to overcome the large intermolecular forces within the ice than it does to simply add kinetic energy in raising the temperature. The latent heats of fusion for some common materials are listed in Table 2.5.1.

For a given mass of a substance:
Heat energy transferred = mass of substance × specific latent heat of fusion

$$Q = mL_{fusion}$$

TABLE 2.5.1 The latent heats of fusion for some common materials

Substance	Melting point (°C)	L_{fusion} (J kg^{-1})
silver	961	0.88×10^5
oxygen	–219	0.14×10^5
lead	327	0.25×10^5
ethanol	–114	1.05×10^5
water	0	3.34×10^5

Worked example 2.5.1

LATENT HEAT OF FUSION

Calculate how much energy must be removed from 2.50 L of water at 0.0°C to produce a block of ice at 0.0°C. Express your answer in kJ.

Thinking	Working
Cooling from liquid to solid involves the latent heat of fusion, as the energy is removed from the water. Calculate the mass of water involved. 1 L of water = 1 kg	2.50 L = 2.50 kg
Use Table 2.5.1 to find the latent heat of fusion for water.	$L_{fusion} = 3.34 \times 10^5$ J kg^{-1}
Use the equation: $Q = mL_{fusion}$	$Q = mL_{fusion}$ $= 2.50 \times 3.34 \times 10^5$ $= 8.35 \times 10^5$ J
Convert to kJ.	840 kJ of energy is removed.

➤ Try yourself 2.5.1

LATENT HEAT OF FUSION

Calculate how much energy must be removed from 5.5 kg of liquid lead at 327°C to produce a block of solid lead at 327°C. Express your answer in kJ.

Latent heat of vaporisation (boiling)

WS 1.1.5

It takes much more energy to convert a liquid to a gas than it does to convert a solid to a liquid. This is because, to convert to a gas, the intermolecular bonds must be broken. During the change of state, the energy supplied is used solely in overcoming the intermolecular bonds. The temperature will not rise until all of the material in the liquid state is converted to a gas, assuming that the liquid is evenly heated. For example, when liquid water is heated to boiling point, a large amount of energy is required to change its state from liquid to steam (gas). The temperature will remain at 100°C until all of the water has turned into steam. Once the water is completely converted to steam, then the temperature can start to rise again.

The amount of energy required to change a liquid to a gas is the same as the potential energy released when the gas returns to a liquid. It is called the **latent heat of vaporisation**.

The amount of energy required will depend on the substance. The energy required, Q, to change a liquid of mass m to a gas can be calculated using the latent heat of vaporisation, L_{vapour}.

$$Q = mL_{vapour}$$

Note that, in just about every case, the latent heat of vaporisation of a substance will be different from the latent heat of fusion for that substance. The latent heats of vaporisation for some common materials are listed in Table 2.5.2.

In many instances, it is necessary to consider the energy required to heat a substance as well as change its state. Problems like this are solved by considering the rise in temperature separately from the change of state.

For a given mass of a substance:

Heat energy transferred = mass of substance × specific latent heat of vaporisation

$$Q = mL_{vapour}$$

TABLE 2.5.2 The latent heat of vaporisation of some common materials

Substance	Boiling point (°C)	L_{vapour} (J kg^{-1})
oxygen	−183	2.2×10^5
ethanol	78	8.7×10^5
lead	1750	9.0×10^5
water	100	22.6×10^5
silver	2193	23.0×10^5

Worked example 2.5.2

CHANGE IN TEMPERATURE AND STATE

50.0 mL of water is heated from a room temperature of 20.0°C to its boiling point at 100.0°C. The water is boiled at this temperature until it has completely evaporated. Calculate how much energy in total was required to raise the temperature and boil the water.

Thinking	Working
Calculate the mass of water involved. 1 L of water = 1 kg	50.0 mL of water = 0.050 kg
Find the specific heat capacity of water from Table 2.4.1.	$c = 4200$ J kg^{-1} K^{-1}
Use the equation $Q = mc\Delta T$ to calculate the heat energy required to change the temperature of water from 20°C to 100°C.	$Q = mc\Delta T$ $= 0.050 \times 4200 \times (100.0 - 20.0)$ $= 16\,800$ J
Find the specific latent heat of vaporisation of water from Table 2.5.1.	$L_{vapour} = 22.6 \times 10^5$ J kg^{-1}
Use the equation $Q = mL_{vapour}$ to calculate the latent heat required to boil water.	$Q = mL_{vapour}$ $= 0.05 \times 22.6 \times 10^5$ $= 113\,000$ J
Find the total energy required to raise the temperature and change the state of the water.	Total $Q = 16\,800 + 113\,000$ $= 1.30 \times 10^5$ J $= 130$ kJ

➤ Try yourself 2.5.2

CHANGE IN TEMPERATURE AND STATE

3.0 L of water is heated from a fridge temperature of 4.0°C to its boiling point at 100.0°C. The water is boiled at this temperature until it has completely evaporated. Calculate how much energy in total was required to raise the temperature and boil the water.

EVAPORATION AND COOLING

If you spill some water on the floor then come back in a couple of hours, the water will probably be gone. It will have evaporated. It has changed from a liquid into a vapour at room temperature in a process called **evaporation**. The reason for this is that the water particles, if they have sufficient energy, can escape through the surface of the liquid into the air. Over time, no liquid remains.

Evaporation is more noticeable in **volatile** liquids such as methylated spirits, mineral turpentine, perfume and liquid paper. The surface bonds are weaker in these liquids, and they evaporate rapidly. This is why you should never leave the lids off bottles of these liquids. Volatile liquids are often stored in narrow-necked bottles for this reason.

The rate of evaporation of a liquid can depend on:

- the volatility of the liquid—more-volatile liquids evaporate faster
- the surface area—greater evaporation occurs when greater surface areas are exposed to the air
- the temperature—hotter liquids evaporate faster
- the humidity—less evaporation occurs in more humid conditions
- air movement—if a breeze is blowing over the surface of the liquid, evaporation is more rapid.

Whenever evaporation occurs, higher-energy particles escape the surface of the liquid, leaving the lower-energy particles behind (Figure 2.5.2). As a result, the average kinetic energy of the particles remaining in the liquid decreases and the temperature drops. Humans and many animals use this cooling principle when sweating (perspiring) to stay cool. Since the hotter molecules of perspiration are removed by evaporation, the overall temperature after evaporation is cooler; it is your perspiration evaporating that cools you down. Similarly, when rubbing alcohol is dabbed on your arm before an injection, the cooling of the volatile liquid numbs your skin.

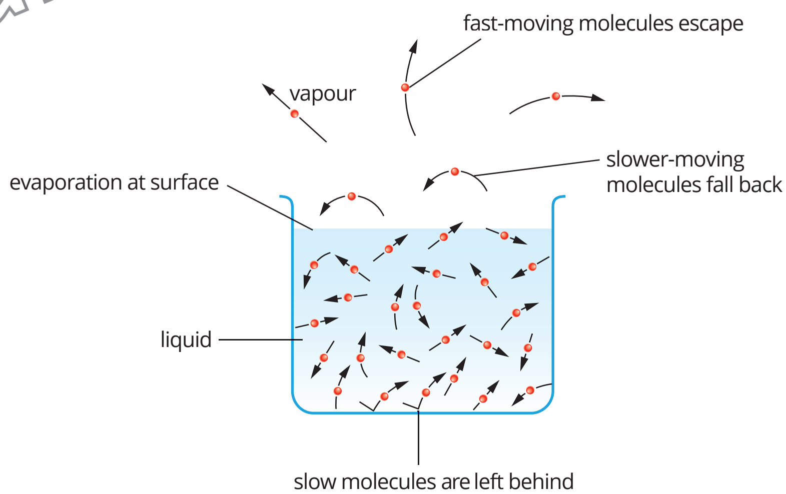


FIGURE 2.5.2 Fast-moving molecules with high kinetic energy can escape the liquid, leaving molecules with lower kinetic energy behind.

HEATING CURVES

Figure 2.5.3 is a heating graph for water. It is also called a phase change diagram. The graph shows the temperature change when a substance is heated at a constant rate and it changes state from a solid to a liquid to a gas. These graphs can provide information on the specific heat capacity of each phase and latent heats of fusion and vaporisation of a substance.

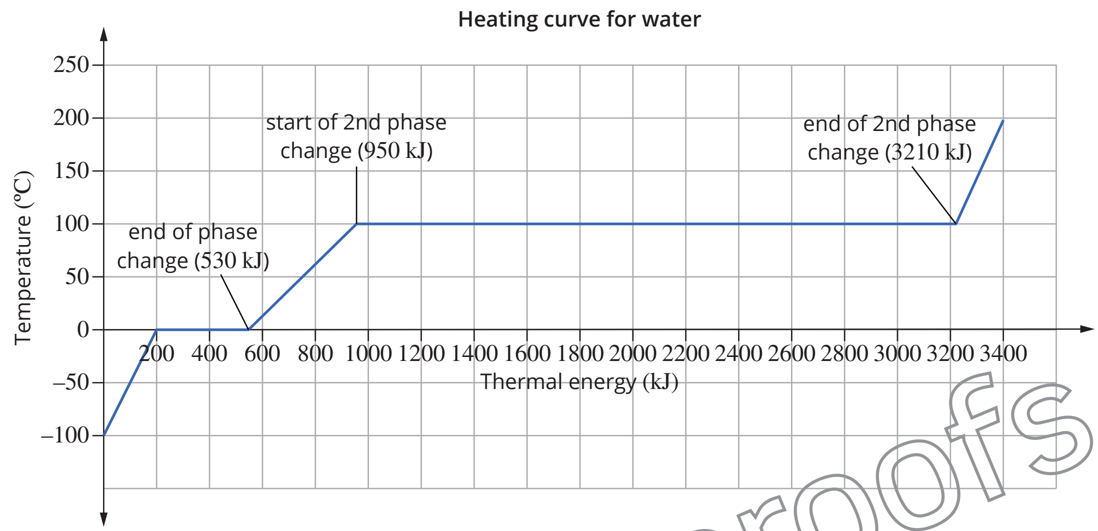


FIGURE 2.5.3 A heating graph for water

Determining latent heats from a heating curve

The horizontal sections of the graph represent the phase changes in the substance. The heat energy going into the substance is being used to increase the potential energy and change phase. By determining how much heat energy (Q) is going into a substance for a phase change to occur, the latent heat for that phase change can be calculated.

Worked example 2.5.3

DETERMINING LATENT HEAT OF FUSION USING A HEATING CURVE

1.0 kg of water has been heated from ice through to steam. Use the appropriate section of Figure 2.5.3 to determine the latent heat of fusion of water.	
Thinking	**Working**
Find the heat required to change the ice to water. On the graph, the first phase change is from 200 KJ to 530 kJ.	$Q = 530\,\text{kJ} - 200\,\text{KJ}$ $= 330\,\text{kJ}$
Rearrange the equation $Q = mL_{\text{fusion}}$ to make L_{fusion} the subject. $Q = 330\,000\,\text{J}$ $m = 1.0\,\text{kg}$	$Q = mL_{\text{fusion}}$ $L_{\text{fusion}} = \frac{Q}{m}$ $= \frac{330000}{1.0}$ $= 330000\,\text{J kg}^{-1}$

➤ Try it yourself 2.5.3

DETERMINING LATENT HEAT OF FUSION USING A HEATING CURVE

1.0 kg of water has been heated from ice through to steam at a temperature of 190°C. Use the appropriate section of Figure 2.5.3 to determine the latent heat of vaporisation of water.

Determining specific heat capacities from a heating curve

The sloped sections of the heating graph indicate when a phase (solid, liquid or gas) is undergoing a temperature change. These are the sections where the specific heat capacity applies. From the gradient of the graph (in these sections), the specific heat capacity can be calculated for each phase.

Worked example 2.5.4

DETERMINING SPECIFIC HEAT CAPACITY USING A HEATING CURVE

1.0 kg of water has been heated from ice through to steam. Use the slope of the appropriate section of Figure 2.5.3 to determine the specific heat capacity of water.

Thinking	Working
Find the gradient of the line for the liquid phase.	Gradient $= \frac{\text{rise}}{\text{run}}$ $= \frac{100}{420\,000}$ $= 23.8 \times 10^{-4}$
For a heating curve, the rise is the change in temperature, ΔT, and the run is the change in energy transferred, ΔQ.	Gradient $= \frac{\text{rise}}{\text{run}}$ $= \frac{T}{Q}$
Rearrange the equation $Q = mc\Delta T$.	$Q = mc\Delta T$ $\frac{T}{Q} = \frac{1}{mc}$
The mass is 1.0 kg, so $m = 1$.	The gradient is $\frac{1}{c}$.
Use the gradient and the result that the gradient is $\frac{1}{c}$ to find c.	$\frac{1}{c} = 23.8 \times 10^{-4}$ $Q = \frac{1}{2.38 \times 10^{-4}}$ $= 4200\ \text{J kg}^{-1}\ \text{K}^{-1}$

➤ Try yourself 2.5.4

DETERMINING SPECIFIC HEAT CAPACITY USING A HEATING CURVE

1.0 kg of water has been heated from ice through to steam. Use the slope of the appropriate section of Figure 2.5.3 to determine the specific heat capacity of steam.

2.5 Review

SUMMARY

- When a solid material changes state, energy is needed to separate the particles by overcoming the attractive forces between the particles.
- Specific latent heat is the energy required to change the state of 1 kg of material at a constant temperature.
- In general, for any mass of material the energy required (or released) is $Q = mL$.
- The specific latent heat of fusion, L_{fusion}, is the energy required to change 1 kg of a material between the solid and liquid states.
- The specific latent heat of vaporisation, L_{vapour}, is the energy required to change 1 kg of a material between the liquid and gaseous states.
- The specific latent heat of fusion of a material will be different from (and usually less than) the latent heat of vaporisation for that material.
- Evaporation is when a liquid turns into gas at room temperature. The temperature of the liquid falls as this occurs.
- The rate of evaporation depends on the volatility, temperature and surface area of the liquid and the presence of air movement.
- A heating curve or phase change diagram shows the change in temperature when a substance is heated at a constant rate and changes from a solid to a liquid to a gas. It can be used to find the latent heat and specific heat capacity of a substance.

KEY QUESTIONS

Refer to the values in Table 2.5.1 and Table 2.5.2. You may also need to refer to Table 2.4.1.

Retrieval

1 State and explain two factors that affect the rate of evaporation of sweat.

2 Define 'latent heat'.

3 What kind of internal energy is associated with a phase change?

Comprehension

4 Explain why the temperature does not change during the boiling phase on a heating curve.

5 Determine the amount of heat energy that must be transferred away from 100.0 g of steam at 100.0°C to change it completely to a liquid (at 100.0°C). ($L_{water\ vapour} = 22.6 \times 10^5$ J kg^{-1}).

6 24 600 J of heat energy is removed from a sample of liquid silver at its melting point to turn it into a solid. How much silver was present in the sample? ($L_{fusion\ silver} = 0.88 \times 10^5$ J kg^{-1}).

Analysis

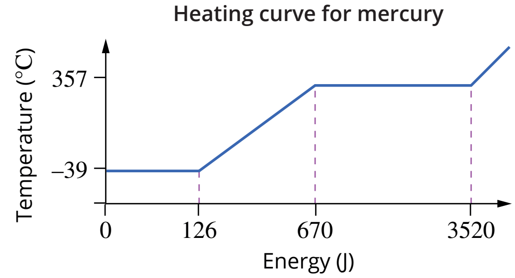
7 How much energy is needed to boil 1.20 kg of ethanol from a starting temperature of 60°C? Assume no loss of energy to the surroundings and that the ethanol completely changes state to a gas. ($L_{ethanol} = 2460 \times 10^5$ J kg^{-1}, $L_{ethanol\ vapour} = 8.7 \times 10^5$ J kg^{-1}).

8 Determine how many kJ of energy are required to melt exactly 100.0 g of ice initially at –4.00°C. Assume no loss of energy to surroundings and that the ice undergoes a full phase change. ($c_{ice} = 2100$ J kg^{-1}K^{-1}, $L_{fusion\ water} = 3.34 \times 10^5$ J kg^{-1}).

9 2.134 × 106 J of energy was used to boil 9.7 kg of a substance. Calculate the latent heat of vaporisation of the substance, then identify the substance using information contained in Table 2.5.2.

10 The graph below represents the heating curve for mercury, a metal that is a liquid at normal room temperature. Thermal energy is added to 10.0 g of solid mercury, initially at a temperature of –39°C, until all of the mercury has evaporated.

Heating curve for mercury



a Explain why the temperature remains constant during the first part of the graph.

b Identify the melting point of mercury, in degrees Celsius.

c Identify the boiling point of mercury, in degrees Celsius.

d Determine the latent heat of fusion of mercury.

e Determine the latent heat of vaporisation of mercury.

11 Draw a heating curve for 300 g of ethanol. Use Tables 2.5.1 and 2.5.2, as well as the specific heat capacities of solid ethanol (970J kg^{-1} K^{-1}), liquid ethanol (2460 J kg^{-1} K^{-1}) and gaseous ethanol (1900 J kg^{-1} K^{-1}). Note that the melting point of ethanol is −114°C and its boiling point is 78°C.

2.6 Calorimetry

BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- understand how heat flows from a hot object to a cold object
- understand thermal equilibrium
- understand the conservation of energy with regard to thermal processes
- solve problems involving specific heat capacity, specific latent heat and thermal equilibrium.

THERMAL EQUILIBRIUM

If you leave a hot cup of coffee on a table, it will cool down until it reaches room temperature. This is because particles in the coffee will 'lose' kinetic energy to the environment until the average kinetic energy of particles in the coffee is the same as that of the environment. Similarly, if you put ice into a warm drink, it will quickly melt to leave the entire drink a little cooler than before the ice was added.

If two objects are in **thermal contact**, energy can flow between them. For example, if an ice cube is placed in a copper pan, the ice molecules are in thermal contact with the copper atoms. Assuming that the copper is warmer than the ice, thermal energy will flow from the copper to the ice. When two objects in thermal contact stop having a flow of energy between them, they are in **thermal equilibrium**. If a frozen piece of steak is placed in a container of warm water, kinetic energy is transferred from the water to the steak. The particles within the steak gain kinetic energy and the steak warms up. The water loses energy and cools down. Eventually the transfer of energy between the steak and water will stop when their particles have the same average kinetic energy. They are now in thermal equilibrium, and the steak and water will be at the same temperature.

At the point of thermal equilibrium, the temperature of the two objects will be the same. The average kinetic energy of the two objects will also be the same because average kinetic energy is a measure of temperature, as shown in Module 2.2. The flow of thermal energy when two objects are in thermal contact is due to a transfer of kinetic energy as molecules collide.

ZEROTH LAW OF THERMODYNAMICS

The topic of thermal physics involves phenomena associated with energy transfer between objects at different temperatures. Since the 19th century, scientists have developed four laws for this subject. The first two will be studied in this chapter.

The first, second and third laws had been known and understood for some time. Then another law was also determined. This final law was so important that it was decided to place it first, and so it is called the zeroth law of **thermodynamics**.

The zeroth law of thermodynamics relates to thermal equilibrium and thermal contact and allows temperature to be defined. Objects in thermal contact with each other will tend towards thermal equilibrium. So, if a thermometer is placed in a glass of water, the thermometer and the water will tend towards thermal equilibrium with each other, allowing the thermometer to measure the temperature of the water. If cordial is added to the glass of water, the water and cordial will come to a thermal equilibrium. The thermometer will also reach a thermal equilibrium with the water and cordial mixture. By this point, the particles within each substance have the same average kinetic energy.

> The zeroth law of thermodynamics:
>
> If objects A and B are each in thermal equilibrium with object C, then objects A and B are in thermal equilibrium with each other.

Heat lost by one substance = heat gained by the other substance

$Q_{lost} = Q_{gained}$

FIGURE 2.6.1 A scientist uses a bomb calorimeter to measure a temperature change to determine the thermal energy transfer.

FIGURE 2.6.2 The Styrofoam cup is a makeshift calorimeter that insulates the process and the thermometer measures the temperature change. The beaker provides stability.

Conservation of energy

In physics, quantities that are conserved are very important. Energy is one of these conserved quantities. Energy cannot be created or destroyed; it simply changes form. In a closed system, a system from which no thermal energy can escape, if two objects are in thermal contact, the thermal energy lost by one object is the thermal energy gained by the other object; that is, $Q_{lost} = Q_{gained}$.

Calorimetry is the activity of measuring an amount of thermal energy. Calorimetry determines the specific heat capacities and latent heats of different substances because of the conservation of energy. To provide a closed system, scientists use insulated instruments called calorimeters to measure amounts of heat (Figure 2.6.1). Calorimeters can be used to measure the energy produced in a chemical reaction or temperature change when two different objects (at different temperatures) are combined. In a high school laboratory, a foam cup can be used to simulate the insulating environment of a calorimeter (Figure 2.6.2). This set-up could be used to determine the specific heat capacity of an unknown metal. The metal can be heated to a known temperature, and then added to some water. The metal and water are allowed to come to equilibrium and the temperature increase of the water will allow the calculation of the specific heat capacity of the metal, as shown in Worked example 2.6.1.

Worked example 2.6.1

DETERMINING SPECIFIC HEAT CAPACITY USING CALORIMETRY

A 20.0 g piece of metal is heated to 100.0°C and then placed in a foam cup containing 50.0 mL of water at 24.0°C. The final temperature of the water and metal cube is 27.0°C. Calculate the specific heat capacity of the metal.

Thinking	Working
Calculate the mass of water involved. 1 L of water = 1 kg	50.0 mL of water = 0.050 kg
Find the specific heat capacity of water from Table 2.4.1 on page XX.	$c = 4200\,\text{J kg}^{-1}\,\text{K}^{-1}$
Use the equation $Q = mc\Delta T$ to calculate the heat energy required to change the temperature of water from 24°C to 27°C.	$Q_{water} = mc\Delta T$ $= 0.05 \times 4200 \times (27 - 24)$ $= 630\,\text{J}$
Recall that, due to conservation of energy, the heat gain by the water will be the heat lost by the metal.	$Q_{metal} = 630\,\text{J}$
Convert the mass of the metal to kilograms.	20.0 g of metal = 0.020 kg
Rearrange $Q_{metal} = mc\Delta T$ to make c the subject.	$Q_{metal} = mc\Delta T$ $c = \frac{Q_{metal}}{m\Delta T}$
Substitute in the values and find the specific heat capacity for a temperature change from 100.0°C to 27.0°C.	$c = \frac{Q_{metal}}{m\Delta T}$ $= \frac{630}{0.02 \times (100 - 27)}$ $= 430\,\text{J kg}^{-1}\,\text{K}^{-1}$

➤ Try yourself 2.6.1

DETERMINING SPECIFIC HEAT CAPACITY USING CALORIMETRY

A 34.6 g piece of metal has been heated to 97.0°C. This is then placed in a Styrofoam cup containing 71 mL of water at 20.0°C. The final temperature of the water and metal cube is 29.5°C. Calculate the specific heat capacity of the metal.

WS 1.1.6

Worked example 2.6.2

DETERMINING FINAL TEMPERATURE OF A MIXTURE USING CALORIMETRY

Four ice cubes (20 g each) at 0.0°C were added to a 250 g glass of water at 30°C. What is the final temperature of the drink?

Thinking	Working
Calculate the mass of the ice and water involved. 1000 g = 1 kg	4 × 20 g ice = 0.08 kg 250 g water = 0.250 kg
Find the specific latent heat of fusion for water from Table 2.5.1.	$L_{fusion} = 3.34 \times 10^5\ \text{J kg}^{-1}$
Find the specific heat capacity of water from Table 2.4.1.	$c = 4200\ \text{J kg}^{-1}\ \text{K}^{-1}$
Recall that, due to the conservation of energy, the heat gained by the ice will be the heat lost by the water.	$Q_{ice} = Q_{water}$ $m_{ice}L_{fusion} + m_{ice}c_{water}\Delta T_{water} = m_{water}c_{water}\Delta T_{water}$ $m_{ice}L_{fusion} + m_{ice}c_{water}(T_{final} - 0) = m_{water}c_{water}(30 - T_{final})$
Rearrange so that all the T_{final} terms are on the left-hand side of the equation.	$m_{ice}L_{fusion} + m_{ice}c_{water}T_{final} - 0m_{ice}c_{water} = 30m_{water}c_{water} - m_{water}c_{water}T_{final}$ $m_{ice}c_{water}T_{final} + m_{ice}c_{water}T_{final} = -m_{ice}L_{fusion} + 30m_{water}c_{water}$ $T_{final}(m_{ice}c_{water} + m_{water}c_{water}) = -m_{ice}L_{fusion} + 30m_{water}c_{water}$
Substitute in values and find T_{final}.	$T_{final}(m_{ice}c_{water} + m_{water}c_{water}) = -m_{ice}L_{fusion} + 30m_{water}c_{water}$ $T_{final}(0.08 \times 4200 + 0.025 \times 4200) = -0.08 \times (3.34 \times 10^5) + 30 \times 0.250 \times 4200$ $T_{final} \times 1386 = 4780$ $T_{final} = 3.4°C$

➤ Try yourself 2.6.2

DETERMINING FINAL TEMPERATURE OF A MIXTURE USING CALORIMETRY

Three ice cubes (15 g each) at 0.0°C were added to a 200 g glass of water at 20°C. What is the final temperature of the drink?

2.6 Review

SUMMARY

- If two objects are in thermal contact, thermal energy will flow from the hot object to the cold object until the two objects are at the same temperature. This is known as thermal equilibrium.
- The zeroth law of thermodynamics states that if objects A and B are each in thermal equilibrium with object C, then objects A and B are in thermal equilibrium with each other. A, B and C must be at the same temperature.
- Conservation of energy states that heat cannot be created or destroyed. This means that the thermal energy lost by one substance is gained by another.
- Calorimetry is the measurement of heat. It can be used to determine the specific heat capacity or latent heat of a substance.

KEY QUESTIONS

Describe

1 Define 'calorimetry'.

2 Define 'thermal equilibrium'.

3 Compare the average kinetic energy of two objects in thermal equilibrium.

Apply

4 Explain why two objects that are in thermal equilibrium with a third object must also be in thermal equilibrium with each other.

5 Explain why putting ice (at 0°C) into a drink is much more effective at cooling your drink down than putting water at 0°C into your drink.

6 What is the difference between adding 50 g of solid ice to a glass of water compared with adding 50 g of crushed ice to a glass of water? In your response, refer to how fast the water cools down and to the final temperature.

Analyse

7 250 mL of water at 30°C is mixed with 1 L of water at 80°C. What is the final temperature of the water?

8 A 1.4 kg piece of metal has been heated to 60.0°C. This is then placed in an insulated tank containing 18.3 kg of water at 29.0°C. The final temperature of the water and metal cube is 29.5°C. Determine the specific heat capacity of the metal and from this deduce what metal you think it is. (c_{water} = 4200 J kg^{-1} K^{-1})

9 A 5.00 kg block of iron at 60.0°C is brought in contact with a 10.0 kg block of aluminium at 30.0°C. Assume no energy is lost to the surroundings. What is the final temperature when the metals reach thermal equilibrium? (c_{iron} = 440 J kg^{-1} K^{-1}, $c_{aluminium\ spirit}$ = 900 J kg^{-1} K^{-1}).

10 An unknown amount of copper at 78.0°C was added to 2.30 kg of methylated spirits at 21.0°C. The final temperature of the mixture was 25.2°C. Calculate the mass of the copper. (c_{copper} = 390 J kg^{-1} K^{-1}, $c_{methylated\ spirit}$ = 2500 J kg^{-1} K^{-1}).

2.7 Mechanical work and efficiency

BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- understand that thermal energy has the capacity to do mechanical work
- understand the conservation of energy and the first law of thermodynamics
- understand efficiency and the loss of useful energy while total energy is conserved
- calculate the efficiency of an energy transfer.

WORK

In the mid-1800s, James Joule, a British brewer, performed a series of experiments that showed the equivalence of heat and mechanical energy (Figure 2.7.1). In his heat equivalency experiment, a mass was dropped, transferring potential energy to kinetic energy (due to the conservation of energy). This dropping mass pulled a cord turning a spindle. The spindle turned some paddles, which heated the liquid, and the temperature change was measured. The kinetic energy of the turning paddles caused the surrounding liquid to increase in temperature. Through this experiment, Joule measured the heat capacity of water with a surprising level of accuracy for the time.

Work is done when energy is transferred and causes an object to move. In Joule's experiment, the turning paddle did work on the liquid, increasing the average kinetic energy of the molecules, thus increasing the temperature of the water. Joule's experiment converted potential energy and then kinetic energy from the falling mass to kinetic energy of the spindle and paddles, and then to thermal energy that heated the liquid.

Any system with thermal energy has the capacity to do mechanical work. For example, heat is created by the chemical reactions in an engine. The heat expands a gas, which moves a piston and work is done.

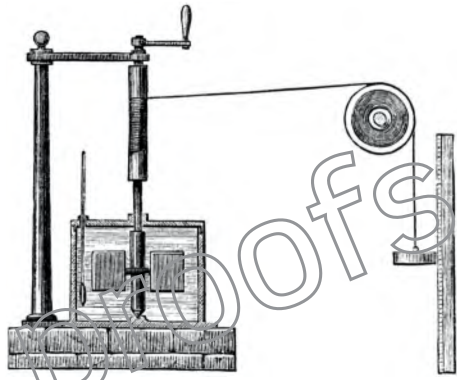


FIGURE 2.7.1 Joule's heat equivalency experiment. James Prescott Joule used this apparatus to determine that mechanical work was equivalent to heat. The mass would drop, pulling the cord and turning the spindle. The turning spindle would turn the paddles, heating the liquid. The change in temperature of the liquid was measured by the thermometer.

THE FIRST LAW OF THERMODYNAMICS

In a closed system, energy is always conserved. Energy cannot be created or destroyed, it can only be transferred or transformed. The first law of thermodynamics states that energy transfers from one form to another and the total internal energy in a system is constant. The internal energy of the system can be changed by heating or cooling, or by work being done on or by the system.

Any change in the internal energy, ΔU, of a system is equal to the energy added by heating, $+Q$, or removed by cooling, $-Q$, plus the work done on, $+W$, or work done by, $-W$, the system. Note that due to the conservation of energy, all energy must be accounted for in the internal energy; there is no 'lost' energy in the system.

$$\Delta U = Q - W$$

The internal energy, U, of a system is defined as the total kinetic and potential energy of the system. As the average kinetic energy of a system is related to its temperature and the potential energy of the system is related to the state, then a change in the internal energy of a system means that either the temperature changes or the state changes.

- If heat, Q, is added to the system, then the internal energy, U, rises by either increasing the temperature or changing state from solid to liquid or liquid to gas. Similarly, if work, W, is done on a system, then the internal energy rises, and the system will once again increase in temperature or change state by melting or boiling. When heat is added to a system or work is done on a system, ΔU is positive.
- If heat, Q, is removed from the system, then the internal energy, U, decreases by either decreasing temperature or changing state from liquid to solid or gas to liquid. Similarly, if work, W, is done by the system, then the internal energy decreases and the system will once again decrease in temperature or change state

> Any change in the internal energy, ΔU, of a system is equal to the energy added by heating, $+Q$, or removed by cooling, $-Q$, minus the work done on, $+W$, or work done by, $-W$, the system.

by condensing or solidifying. When heat is removed from a system or work is done by a system, ΔU is negative.

- If heat is added to the system and work is done by the system, then the magnitude of the energy into the system compared to the magnitude of the energy out of the system determines whether the internal energy increases (ΔU is positive) or decreases (ΔU is negative).

Worked example 2.7.1

CALCULATING THE CHANGE IN INTERNAL ENERGY

A 1 L beaker of water has 25 kJ of work done on it and loses 30 kJ of thermal energy to the surroundings. Calculate the change in energy of the water.

Thinking	Working
Heat is removed from the system, so Q is negative. Work is done on the system, so W is positive.	$\Delta U = Q + W$ $= -30 - (-25)$
Note that the units are kJ, so express the final answer in kJ.	$= -5$ kJ

➤ Try yourself 2.7.1

CALCULATING THE CHANGE IN INTERNAL ENERGY

A student places a heating element and a paddle-wheel apparatus in an insulated container of water. She calculates that the heater transfers 2530 J of thermal energy to the water and the paddle does 240 J of work on the water. Calculate the change in internal energy of the water.

EFFICIENCY

Energy can never be created or destroyed; however, it may transform into less useful forms of energy. In any heat transfer in the real world, some of the heat is 'wasted' to the surroundings. This wasted heat reduces the useable energy. The percentage of energy that is effectively transformed in an energy transfer process is called its efficiency. For example, an incandescent light bulb is only 10% energy efficient; that is 10% of the energy output is light and 90% is wasted as heat. In comparison, a LED bulb is 90% energy efficient. There is now much more effort going into designing energy-efficient buildings and devices than in the past.

The efficiency of transferring energy from one form to another is expressed as:

$$\text{Efficiency } (\eta) = \frac{\text{useful energy transferred}}{\text{total energy}} \times \frac{100}{1}\%$$

$$\eta = \frac{\text{energy output}}{\text{energy input}} \times \frac{100}{1}\%$$

$\eta = \frac{\text{energy output}}{\text{energy input}} \times \frac{100}{1}\%$

Worked example 2.7.2

CALCULATING EFFICIENCY

A particular Bunsen burner provides 76 000 J of heat to a beaker of water. The useful energy that heated the water is 27 000 J. The rest of the heat is lost to the environment. Calculate the efficiency of this process.

Thinking	Working
Recall the equation for efficiency.	energy output = 27 000 J
Substitute the given values into the equation and calculate η.	energy input = 76 000 J $\eta = \frac{\text{energy output}}{\text{energy input}} \times \frac{100}{1}\%$ $= \frac{27000}{76000} \times \frac{100}{1}\%$ $= 36\%$

➤ Try yourself 2.7.2

CALCULATING EFFICIENCY

A particular barbeque produces 295 000 J of energy but the heat absorbed by the food is just 81 000 J. Determine the efficiency of this process.

2.7 Review

SUMMARY

- Thermal energy has the capacity to do mechanical work.
- The first law of thermodynamics states that energy simply changes from one form to another and the total energy in a system is constant.
- Any change in the internal energy (ΔU) of a system is equal to the energy added by heating ($+Q$) or removed by cooling ($-Q$), minus the work done on ($-W$) or by ($+W$) the system: $\Delta U = Q - W$.
- The efficiency, η, of an energy transfer from one form to another is given by η:

$$\eta = \frac{\text{energy output}}{\text{energy input}} \times \frac{100}{1}\%$$

KEY QUESTIONS

Describe

1. Define 'conservation of energy'.
2. Explain whether it is possible to have greater than 100% efficiency in a thermal energy transfer.
3. Considering conservation of energy, would beating eggs increase or decrease their temperature?

Apply

4. Calculate the efficiency of a power station that transforms 1.4×10^7 J of heat energy into 5.0 MJ of electrical energy.
5. Calculate the efficiency of a hair dryer with an input energy of 480 kJ and a useful output energy of 326 kJ.
6. New household solar panels have an efficiency of 22%. If 82 MJ of solar energy is shining on the solar panel in a day:
   a. how much electrical energy is being produced?
   b. how much energy is being wasted?
7. Electric cars have an efficiency of power to the wheels of 77%. Petrol cars have an efficiency of power to the wheels of 25%. If a car needs 19 kJ per second to drive, calculate:
   a. how much input energy the electric car requires.
   b. how much input energy the petrol car requires.

Analyse

8. What is the efficiency of a kettle when an input energy of 584 kJ is used to heat 1.20 kg of water to raise its temperature from 23.0°C to 100°C? (c_{water} = 4200 J kg^{-1} K^{-1})
9. A chef vigorously stirs a pot of cold water and does 150 J of work on the water. The water also gains 75 J of thermal energy from the surroundings. Calculate the change in energy of the water.
10. A scientist very carefully does mechanical work on a container of liquid sodium. The liquid sodium loses 300.0 J of energy to its surroundings but gains 250.0 J of energy overall. Calculate how much work the scientist did.

Chapter review

02

KEY TERMS

absolute zero
absorption
calorimetry
conduction
conductor
convection
electromagnetic spectrum
emission
evaporation
frequency
heat
incident
insulator
internal energy
Kelvin
kinetic energy
kinetic particle model
latent heat
latent heat of fusion
latent heat of vaporisation
model
Maxwell–Boltzmann distribution
potential energy
radiation
specific heat capacity
temperature
thermal contact
thermal energy
thermal equilibrium
thermodynamics
volatile
wavelength
work

KEY QUESTIONS

Retrieval

1 Identify which of the following statements can be said to describe particles of matter, according to the kinetic particle model.

A They are in constant motion at any temperature.
B They are stationary at all temperatures.
C They are in motion above a certain temperature.
D none of these

2 Which of the following will affect the rate at which an object radiates thermal energy?

A its temperature
B its colour
C its surface nature (shiny or dull)
D all of A–C

3 Identify the units for specific heat capacity.

A J kg K^{-1}
B J K
C m J K^{-1}
D J kg-1 K^{-1}

4 Consider a solid at melting point undergoing a phase change to a liquid. During that phase change, what can be said of the internal energy and the potential energy of the substance?

	Internal energy	Potential energy
A.	Increases	Increases
B.	Increases	Remains the same
C.	Decreases	Remains the same
D.	Remains the same	Increases

5 State what the first law of thermodynamics says about the total internal energy of a closed system.

6 A heating curve is a graph of *heat* (*x*-axis) against *change in temperature* (*y*-axis). What does the gradient of the liquid phase of a heating curve represent?

Comprehension

7 **a** An uncooked chicken is placed into an oven that has been preheated to 180°C. Which of the following statements describes what happens as soon as the chicken is placed in the oven? You can choose more than one.

A Thermal energy flows from the chicken into the hot air.
B The chicken and the air in the oven are in thermal equilibrium.
C Thermal energy flows from the hot air into the chicken.
D The chicken and the air in the oven are not in thermal equilibrium.

b A chicken is inside an oven that has been preheated to 180°C. The chicken has been cooking for one hour and its temperature is also 180°C. Identify which of the following statements best describes this scenario.

A Thermal energy flows from the chicken into the hot air.
B The chicken and the air in the oven are in thermal equilibrium.
C Thermal energy flows from the hot air into the chicken.
D The chicken and the air in the oven are not in thermal equilibrium

8 Identify which of the following temperatures cannot possibly exist. You can choose more than one.
A 1 000 000°C
B –50°C
C –50 K
D –300°C

9 Two cubes, one of silver and one of iron, have the same mass and temperature. A quantity of heat, Q, is removed from each cube. Identify which one of the following properties causes the final temperatures of the cubes to be different.
A density
B specific heat capacity
C latent heat of vaporisation
D volume

10 Consider two objects that are in thermal contact in a closed system. Explain the relationship between the heat lost from one object and the heat gained by the second object.

11 Explain how temperature differs from heat.

12 Convert:
a 5.0°C to kelvin
b 200.0 K to °C.

13 Explain why metals are more likely than wood to conduct heat.

14 A solid substance is heated but its temperature does not change. Explain what is occurring.

15 Calculate how many kilojoules of energy are required to melt exactly 80.0 g of silver ($L_{fusion} = 0.88 \times 10^5$ J kg^{-1}).

16 What are the units of the ratio $\frac{\text{specific latent heat of vaporisation}}{\text{specific heat capacity}}$?

17 The heating curve for a solid shows how temperature (vertical axis) varies with heat (horizontal axis). The curve has slope G. Write an equation to calculate the mass of the solid based on the gradient.

18 Two 20.0 g cubes of ice (at –4.00°C) are added to a 20.0°C cup of water (250.0 g). Determine the final temperature of the cup of water. (c_{ice} = 2100 J kg^{-1} K^{-1}, c_{water} = 4200 J kg^{-1} K^{-1}, $L_{fusion\ water}$ = 3.34 × 10^5 J kg^{-1})

19 Based on the kinetic particle model, if a substance has no thermal contact with its surroundings, does the temperature of the substance change over time?

Analysis

20 Three samples of a liquid are added to a foam cup: 20 mL at 45°C, 30 mL at 20°C and 10 mL at 0°C. What is the equilibrium temperature of the liquid?
A 15°C
B 20°C
C 25°C
D 30°C

21 Object A is heated from 25°C to 40°C. Object B is heated from 0°C to 15°C. What is the difference in heat energy provided to both objects?
A. There is no difference.
B. More heat is transferred to Object B.
C. Less heat is transferred to Object B.
D. The answer cannot be determined with the information provided.

22 A mass of water is in a calorimeter. A piece of metal half the mass of the water is placed in the water. The water changes temperature by 5 K while the metal has changed temperature by 70 K.
What is the ratio $\frac{\text{specific heat capacity of the water}}{\text{specific heat capacity of the metal}}$?
A 5
B 7
C 9
D 10

23 A sealed, thermally insulated container contains ice at 0°C and water at 15°C. Describe what happens to the total internal energy of the system over time.
A It increases.
B It decreases.
C It increases until the ice melts, then stays the same.
D It remains constant.

24 Explain the difference between internal energy and thermal energy.

25 A solid and a liquid are at thermal equilibrium with each other. Is the thermal energy and/or internal energy of the solid the same as that of the liquid?

26 Explain the difference between solids, liquids and gases using the kinetic theory.

27 A body at a uniform temperature of 300.0 K and an internal energy of 6.0 × 10^6 J is split into two equal halves.
a Has any heat been exchanged?
b What is the temperature of each half?
c Determine the internal energy of each half.

28 Calculate how much energy, in J, is needed to raise the temperature of a 1.0 kg block of aluminium by 20.0°C. ($c_{aluminium}$ = 900 J kg^{-1} K^{-1}).

29 A 2.00 kg metal object requires 5.02×10^3 J of heat to raise its temperature from 20.0°C to 30.0°C. Determine the specific heat capacity of the metal in J kg^{-1} K^{-1}. Give your answer to the nearest whole number.

30 If 48 000 J of energy is given off when a 2.0 kg sample of a liquid cools by 12 K, determine the specific heat capacity of the liquid.

31 An unknown quantity of ice at 0.0°C is added to a 750 g sample of water, the initial temperature of which is 15.0°C. The water cools down to its freezing point. How much ice was added? (c_{ice} = 2100 J kg^{-1} K^{-1}, c_{water} = 4200 J kg^{-1} K^{-1}, $L_{fusion\ water}$ = 3.34×10^5 J kg^{-1}).

32 Hypothermia is the cooling of the body to levels considerably lower than normal. The body's functions slow and death can result. A person may survive 12 hours in air at 0°C before suffering hypothermia but may survive only a few minutes in water at 0°C. Consider why this is so and how wetsuits can help you survive in cold water.

33 You have 30.0 g of ethanol at 22°C, which you pour into 90.0 g of water at 80.0°C. Assuming no heat is produced or lost during this process, calculate the final temperature of the mixture. ($c_{ethanol}$ = 2460 J kg^{-1} K^{-1}, c_{water} = 4200 J kg^{-1} K^{-1}).

34 A heater that is 25% efficient receives 3000.0 J per second. Determine how much of this energy is put out as heat.

35 Describe a system in which thermal energy is converted into mechanical work.

Knowledge utilisation

36 Create an analogy for the three main states of matter (solid, liquid and gas) to explain them to a younger student.

37 Design a cold storage room, powered by a small refrigeration unit, for keeping food fresh in a hot climate. Draw a sketch of your design, clearly showing its insulating features and materials and describe how it minimises all forms of heat transfer.

38 You have a small cube of an unknown metal that you want to identify. Your teacher has suggested that measuring its specific heat capacity would help you identify it. Design an experiment to measure the specific heat capacity of the cube.

Data analysis

An experiment was conducted to determine the specific heat capacity of a metal cube and thus identify the metal. The metal cube was placed in near-boiling water to bring it up to 100°C. The metal cube was then transferred to a known mass of water at room temperature in a foam cup. The temperature change of the water in the foam cup was measured and can be seen in Figure 1. A list of specific heat capacities of common metals is provided in Table 1.

The mass of the metal cube was 11.2 g and it had an initial temperature in the near-boiling water of 95.4°C.

The mass of the room-temperature water was 67.1 g.

The specific heat capacity of water is 4200 J kg^{-1} K^{-1}.

FIGURE 1 Water temperature in the foam cup when the metal cube is added

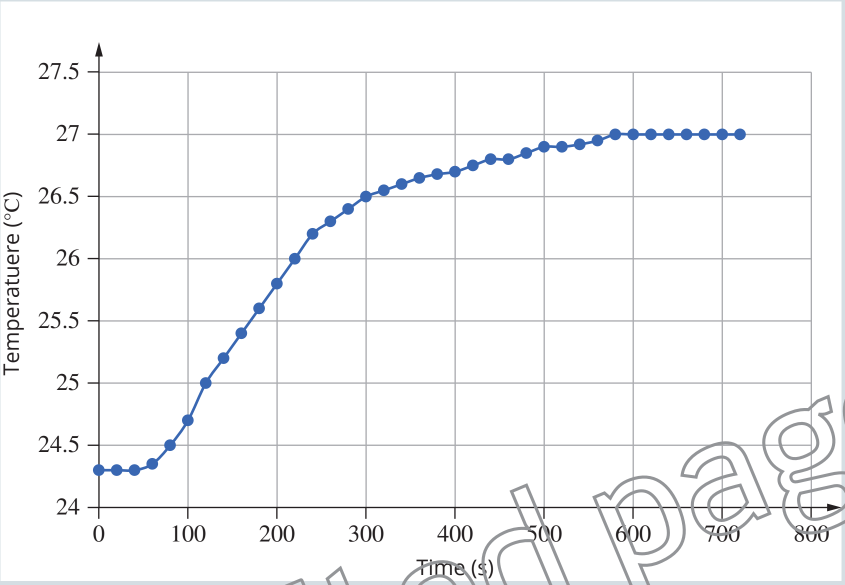


TABLE 1: Specific heat capacities of common metals

Metal	C (J $kg^{-1}K^{-1}$)
lead	130
brass	370
copper	390
iron	440
stainless steel	510
aluminium	900

Question 1 *(1 mark)*

What was the final temperature of the water in the foam cup?

Question 2 *(1 mark)*

What is the temperature difference experienced by the metal cube?

Question 3 *(4 marks)*

Determine the specific heat capacity of the metal.

Question 4 *(1 mark)*

Identify the metal.

Question 5 *(2 marks)*

If a small amount of near-boiling water was mistakenly added along with the metal cube to the room temperature water, explain what effect this would have on the specific heat capacity calculated.

Question 6 *(2 marks)*

If heat is lost through the foam cup, explain what effect this would have on the calculated specific heat capacity. Why is this form of heat lost not considered a large uncertainty for this experiment?