Heinemann

PHYSICS 11

4TH EDITION

VCE Units 1 & 2

Written for the VCE Physics Study Design 2016–2021

ALWAYS LEARNING

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Heinemann PHYSICS 11

4TH EDITION

VCE Units 1 & 2

COORDINATING AUTHOR:
Greg Moran

AUTHORS:
Doug Bail
Keith Burrows
Rob Chapman
Ann Conibear
Carmel Fry
David Madden
John Nicholson
Jeffrey Stanger

CONTRIBUTING AUTHORS:
Daniela Nardelli
Paul Cuthbert
Mark Hamilton
Alistair Harkness
Jack Jurica
Craig Tilley

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Series Consultant: Malcolm Parsons
Publishers: Cameron Salton and Rachel Davis
Project Managers: Jane Sunderland and Hannah Turner
Lead Development Editor: Vicky Chadfield
Editors: Samantha Trafford, Nick Tapp and Marta Veroni
Designer and Cover Designer: Anne Donald
Typesetter: Cam McPhail
Copyright and Picture Editors: Samantha Russell-Tulip
Desktop Operator: Rob Curulli
Illustrators: Pearson India, Guy Holt, Margaret Hastie, Brent Hagan, Chris Hurley, Pat Kermode, Cynthia Nge, Wendy Gorton and Bruce Rankin

Physics 11 Lightbook
Project lead: Catriona McKenzie
Publishers: Pamela Ferrada, Laura Tubino and Monica Schaak
Project Managers: Jade Worcester and Tamara Pirois
Development Editor: Brigid Brignell
Copyright and Picture Editors: Jenny Jones
Lead Editor: Mahtab Dubash
Digital Editor: Averil Robertson
Editors: LearningMate

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4TH EDITION

Writing and Development Team

We are grateful to the following people for their time and expertise in contributing to the Heinemann Physics 11 project.

Malcolm Parsons
Publishing Consultant
Series Consultant

Greg Moran
Head of Department
Coordinating Author

Daniela Nardelli
Learning Development Leader
Subject Lead and Author (Print)

Craig Tilley
Science Writer and Assessor
Associate Subject Lead (Print)

Doug Bail
Education Consultant
Author

Rachael Barker
Teacher
Reviewer

Keith Burrows
Educator
Author

Rob Chapman
Educator
Author

Ann Conibear
Teacher
Author

Paul Cuthbert
Teacher
Author

Carmel Fry
Head of Senior Years Curriculum
Author

Mark Hamilton
Teacher
Author

Alistair Harkness
Teacher
Author

Matt Jones
Teacher
Reviewer

Jack Jurica
Teacher
Author

Melanie Lane
Head of Science
Author

David Madden
Acting Manager, QCAA
Author

Svetlana Marchouba
Laboratory Technician
Author

Dr John Nicholson
Head of Department
Author

Jeffrey Stanger
Head of Department
Author

Contents

Unit 1: What ideas explain the physical world?

AREA OF STUDY 1
How can thermal effects be explained?

Chapter 1 **Heating processes** 1
1.1 Heat and temperature 2
1.2 Specific heat capacity 10
1.3 Latent heat 13
1.4 Conduction 18
1.5 Convection 22
1.6 Radiation 25
Chapter 1 Review 28

Chapter 2 **Applying thermodynamic principles** 29
2.1 Heating by radiation 30
2.2 The enhanced greenhouse effect 40
2.3 Scientific modelling: The enhanced greenhouse effect 55
2.4 Issues related to thermodynamics 64
Chapter 2 Review 81

Area of Study 1 Review 83

AREA OF STUDY 2
How do electric circuits work?

Chapter 3 **Electrical physics** 87
3.1 Behaviour of charged particles 88
3.2 Electric current and circuits 95
3.3 Energy in electric circuits 102
3.4 Resistance 110
Chapter 3 Review 121

Chapter 4 **Practical electric circuits** 123
4.1 Series and parallel circuits 124
4.2 Using electricity 138
4.3 Electrical safety 150
Chapter 4 Review 158

Area of Study 2 Review 161

AREA OF STUDY 3
What is matter and how is it formed

Chapter 5 **The origins of everything** 167
5.1 Measurements in the universe 168
5.2 The big bang 178
5.3 Particles of the Standard Model 186
Chapter 5 Review 198

Chapter 6 **Particles in the nucleus** 199
6.1 Atoms, isotopes and radioisotopes 200
6.2 Radioactivity 205
6.3 Properties of alpha, beta and gamma radiation 213
6.4 Half-life and decay series 218
Chapter 6 Review 223

Chapter 7 **Energy from the atom** 225
7.1 Nuclear fission and energy 226
7.2 Nuclear fusion 234
7.3 Electromagnetic waves and synchrotron radiation 238
7.4 The production of light 245
Chapter 7 Review 250

Area of Study 3 Review 251

Unit 2: What do experiments reveal about the physical world?

AREA OF STUDY 1
How can motion be described and explained?

Chapter 8 **Scalars and vectors** 257
8.1 Scalars and vectors 258
8.2 Adding vectors in one and two dimensions 265
8.3 Subtracting vectors in one and two dimensions 272
8.4 Vector components 279
8.5 Mass and weight 282
Chapter 8 Review 286

Chapter 9 **Linear motion** 287
9.1 Displacement, speed and velocity 288
9.2 Acceleration 300
9.3 Graphing position, velocity and acceleration over time 305
9.4 Equations for uniform acceleration 315
9.5 Vertical motion 321
Chapter 9 Review 327

Chapter 10 **Momentum and force** 331
10.1 Newton's first law 332
10.2 Newton's second law 340
10.3 Newton's third law 348
10.4 Momentum and conservation of momentum 354
10.5 Momentum transfer 362
10.6 Momentum and net force 366
Chapter 10 Review 374

Chapter 11 **Equilibrium of forces** 375
11.1 Torque 376
11.2 Translational equilibrium 388
11.3 Static equilibrium 394
Chapter 11 Review 407

Chapter 12 **Energy, work and power** 411
12.1 Work 412
12.2 Mechanical energy 421
12.3 Using energy: Power and efficiency 433
Chapter 12 Review 443

Area of Study 1 Review 445

AREA OF STUDY 2
Options

Chapter 13 **Stars** 453
13.1 Astronomical measurements 454
13.2 Classifying stars 465
13.3 The life and death of stars 473
Chapter 13 Review 487

Chapter 14 **Forces in the human body** 489
14.1 External forces acting on human body 490
14.2 Forces cause rotation 499
14.3 Tissue under load: Stress and strength 507
14.4 Properties of human tissue 512
14.5 The future: Materials for use in prosthetics 524
Chapter 14 Review 528

Chapter 15 **Energy from nuclear power** 531
15.1 Energy from the nucleus 532
15.2 Nuclear energy as a power source 544
Chapter 15 Review 551

Chapter 16 **Nuclear medicine** 553
16.1 Producing medical radiation 554
16.2 Measurement of radiation doses 563
16.3 Radiation in diagnosis and treatment of human disease 569
Chapter 16 Review 576

Chapter 17 **Particle accelerators** 577
17.1 Synchrotrons 578
17.2 Colliders and particle physics 585
17.3 The importance of accelerator technology in society 596
Chapter 17 Review 602

Chapter 18 **Sport** 603
18.1 Collisions 604
18.2 Sliding and rolling 699
18.3 Hitting, kicking or throwing 616
18.4 The flight of a ball 623
18.5 Air resistance 631
Chapter 18 Review 636

Area of Study 2 Review 638

AREA OF STUDY 3
Practical investigation

Chapter 19 **Practical investigation** 645
19.1 Designing and planning the investigation 648
19.2 Conducting investigations and recording and presenting data 656
19.3 Discussing investigations and drawing evidence-based conclusions 662
Chapter 19 Review 668

APPENDIX 669
ANSWERS 682
GLOSSARY 695
INDEX 702

How to use this book

Heinemann Physics 11 4th edition

Heinemann Physics 11 4th edition has been written to the new VCE Physics Study Design 2016 – 2021. The book covers Units 1 and 2 in an easy-to-use resource. Explore how to use this book below.

Extension

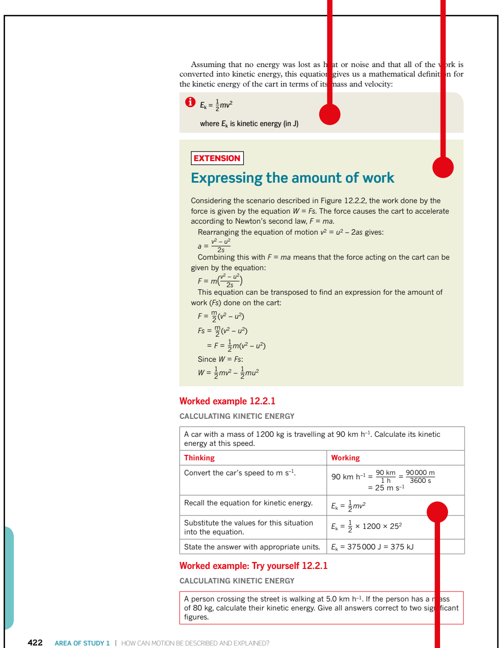
Extension material goes beyond the core content of the Study Design. It is intended for students who wish to expand their depth of understanding.

Highlight

Focus on important information such as key definitions, formulae and summary points.

Chapter opener

Chapter opening pages links the Study Design to the chapter content. Key knowledge addressed in the chapter is clearly listed.



Physics in Action

Physics in Action place physics in an applied situation or relevant context. These refer to the nature and practice of physics, applications of physics and the associated issues and the historical development of concepts and ideas.

PhysicsFile

PhysicsFiles include a range of interesting information and real world examples.

Worked examples

Worked examples are set out in steps that show both thinking and working. This enhances student understanding by linking underlying logic to the relevant calculations.

Each Worked example is followed by a Try Yourself: Worked example. This mirror problem allows students to immediately test their understanding.

Fully worked solutions to all Try Yourself: Worked examples are available on *Heinemann Physics 11 4th edition ProductLink*.

Section summary

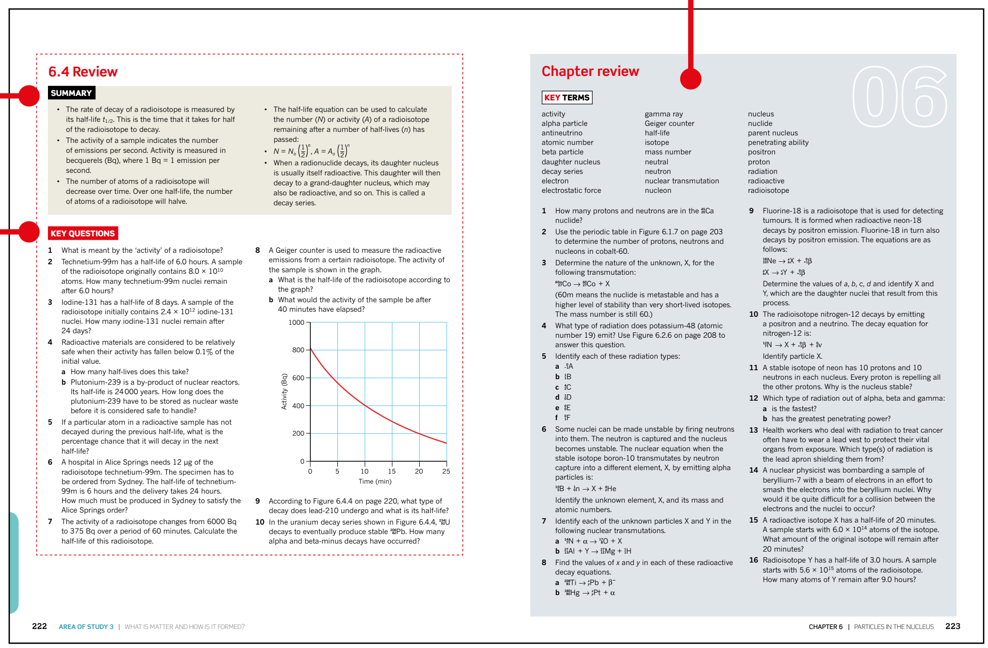
Each section includes a summary to assist students consolidate key points and concepts.

Section review

Each section finishes with questions to test students' understanding and ability to recall the key concepts of the section.

Chapter review

Each chapter finishes with a set of higher order questions to test students' ability to apply the knowledge gained from the chapter.



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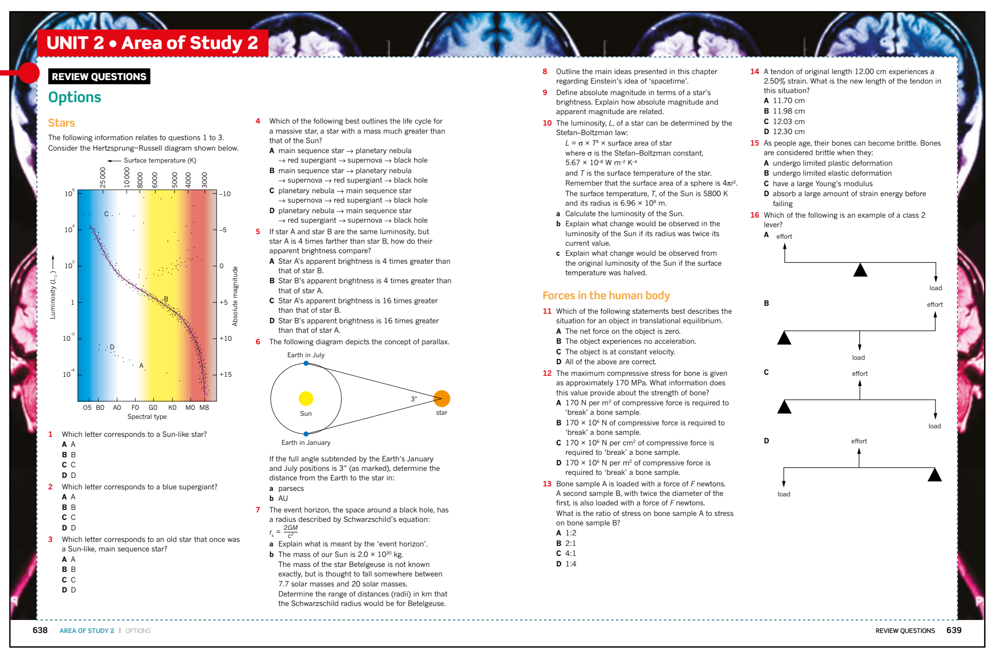

Answers

Numerical answers and key short response answers are included at the back of the book. Comprehensive answers and fully worked solutions for all section review questions, Try Yourself: Worked examples, chapter review questions and Area of Study review questions are provided via *Heinemann Physics 11 4th edition ProductLink.*

Glossary

Key terms are shown in bold and listed at the end of each chapter. A comprehensive glossary at the end of the book includes and defines all key terms.

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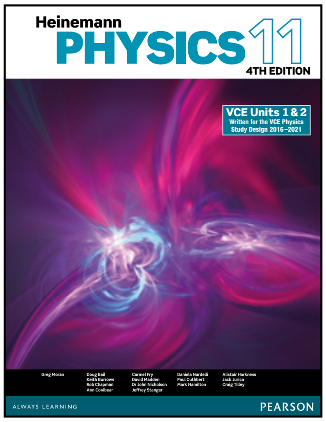

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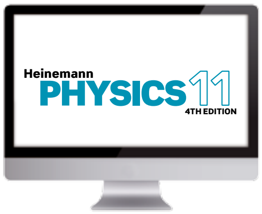

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UNIT 1 What ideas explain the physical world?

AREA OF STUDY 1

How can thermal effects be explained?

Outcome 1: On completion of this unit the student should be able to apply thermodynamic principles to analyse, interpret and explain changes in thermal energy in selected contexts, and describe the environmental impact of human activities with reference to thermal effects and climate science concepts.

AREA OF STUDY 2

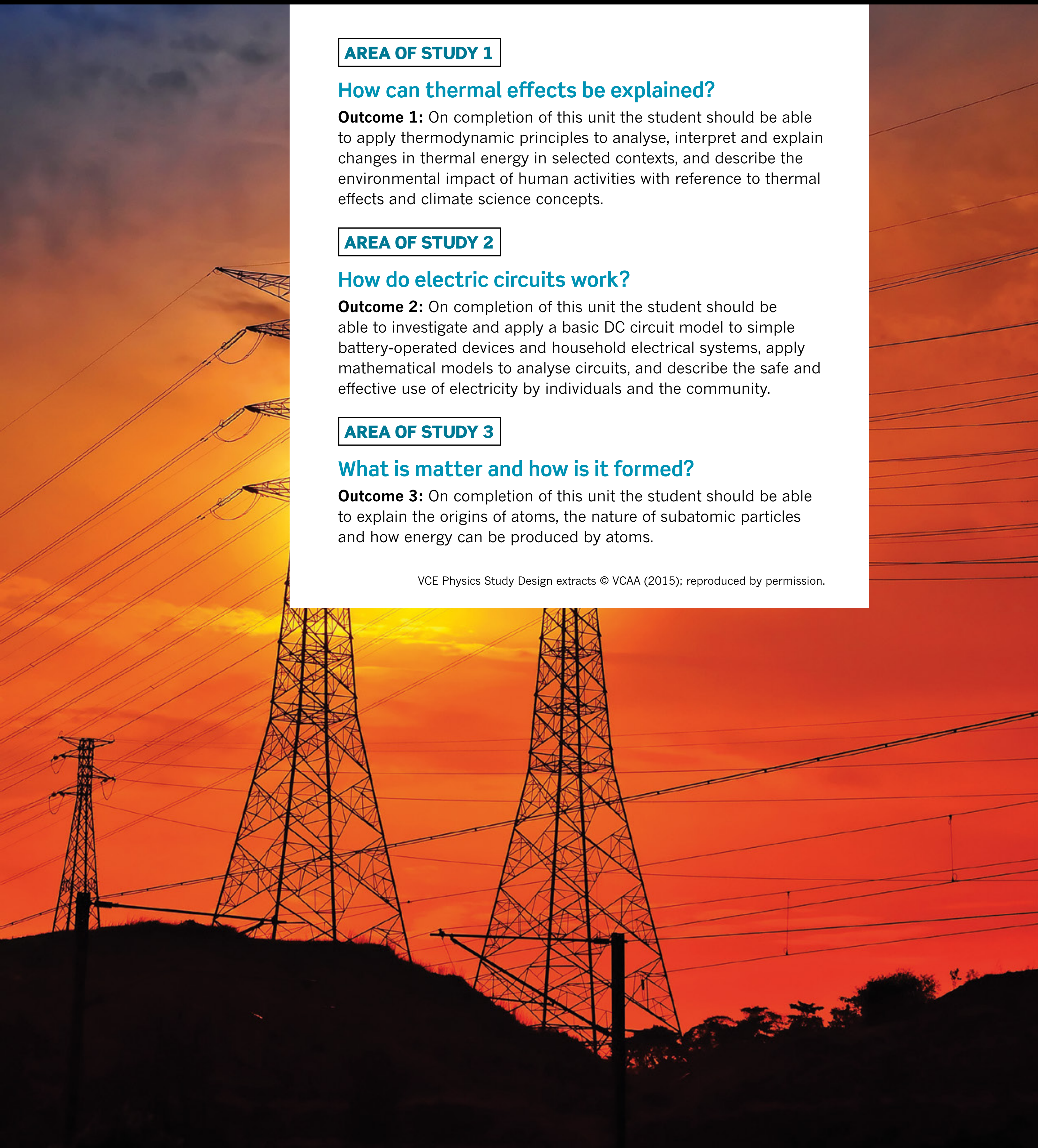
How do electric circuits work?

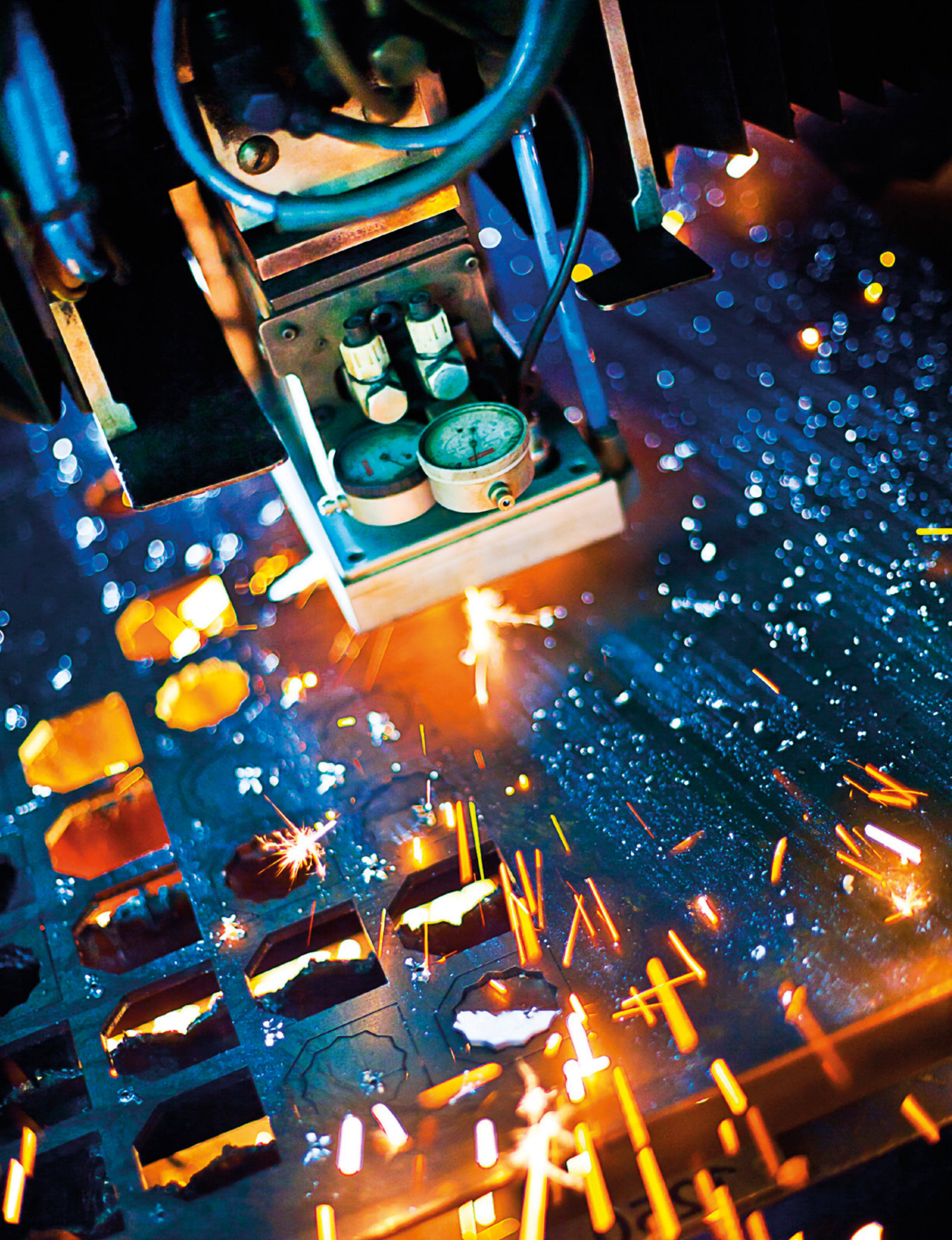
Outcome 2: On completion of this unit the student should be able to investigate and apply a basic DC circuit model to simple battery-operated devices and household electrical systems, apply mathematical models to analyse circuits, and describe the safe and effective use of electricity by individuals and the community.

AREA OF STUDY 3

What is matter and how is it formed?

Outcome 3: On completion of this unit the student should be able to explain the origins of atoms, the nature of subatomic particles and how energy can be produced by atoms.





CHAPTER 01

Heating processes

Due to increasing levels of carbon dioxide in the atmosphere, the Earth is getting warmer. The last two decades of the twentieth century were the warmest for over 400 years. In 2014, the Earth was the warmest since records began in 1890. These higher temperatures increase the severity of bushfires. The rate of evaporation from pastures also increases, drying the land and reducing the level of food production.

Thermal energy is part of our everyday experience. Humans can thrive in the climatic extremes of the Earth, from the outback deserts to ski slopes in winter.

Key knowledge

By the end of this chapter, you will have covered material from the study of the thermodynamic principles related to heating processes, including concepts of temperature, energy and work, and will be able to:

- convert temperature between degrees Celsius and kelvin
- describe the zeroth law of thermodynamics as two bodies in contact with each other coming to a thermal equilibrium
- describe temperature with reference to the average kinetic energy of the atoms and molecules within a system
- investigate and apply theoretically and practically the first law of thermodynamics to simple situations: $Q = U + W$
- explain internal energy as the energy associated with random disordered motion of molecules
- distinguish between conduction, convection and radiation with reference to heat transfer within and between systems
- investigate and analyse theoretically and practically the energy required to:
  - raise the temperature of a substance: $Q = mc\Delta T$
  - Change the state of a substance: $Q = mL$
- Explain why cooling results from evaporation using a simple kinetic energy model.

1.1 Heat and temperature

In the sixteenth 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 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 this section.

KINETIC PARTICLE MODEL

Some philosophers of the Middle Ages 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** (kinetic theory).

Recall that a model is a representation that describes or explains the workings within an object, system or idea.

These are the assumptions behind 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 and gases.

Solids

Within a solid, the particles must be exerting attractive forces or bonds on each other for the matter to hold together in its fixed shape. There must also be 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 (see Figure 1.1.1(a)). But 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.

Liquids

Within a liquid, there is still a balance of attractive and repulsive forces. 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 (see Figure 1.1.1(b)). Generally, the liquid takes up a slightly greater volume than it would in the solid state. Particles collide but remain attracted to each other, so the liquid remains within a fixed volume but with no fixed shape.

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 (see Figure 1.1.1(c)). 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.

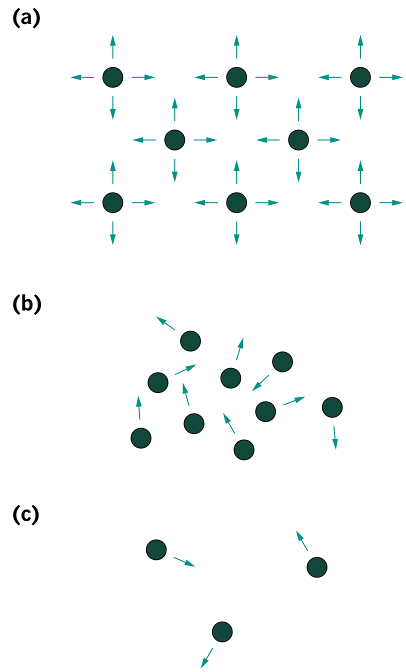


FIGURE 1.1.1 (a) Molecules in a solid have low kinetic energy and vibrate around average positions within a regular arrangement.
(b) The particles in a liquid have more kinetic energy than those in a solid. They move more freely and will take the shape of the container.
(c) Gas molecules are free to move in any direction.

PHYSICSFILE

States of matter: Plasma

The three basic phases (states) of matter are solid, liquid and gas. These are generally all that are discussed in secondary science.

There are in fact four phases of matter that are observable in everyday life—solid, liquid, gas and plasma. Under special conditions, several more exist. 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 1.1.2).

FIGURE 1.1.2 99.9 per cent of the visible universe is made up of plasma.

THE KINETIC PARTICLE MODEL, INTERNAL ENERGY AND TEMPERATURE

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 body to a colder one. Heating is observed by the change in **temperature**, the change of state or the expansion of a substance.

When a solid substance is 'heated', the particles within the material gain either **kinetic energy** (move faster) or **potential energy** (move away from their equilibrium positions).

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. The movement of the particles in a substance due to kinetic energy is ordered: the particles move back and forth and we can model their behaviour. In comparison, the internal energy of a system is associated with the chaotic motion of the particles—it concerns the behaviour of a large number of particles that all have their own kinetic and potential energy.

PHYSICS IN ACTION

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.

Energy is a measure of an object's ability to do **work**. For example, raising an object's temperature or lifting an object is referred to as doing work. Work is measured in joules. The symbol for joules is J. Figure 1.1.3 shows the amount of joules available from some energy sources.

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. Potential energy is stored energy. There are many forms of potential energy, for example gravitational, nuclear, spring and chemical. 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.

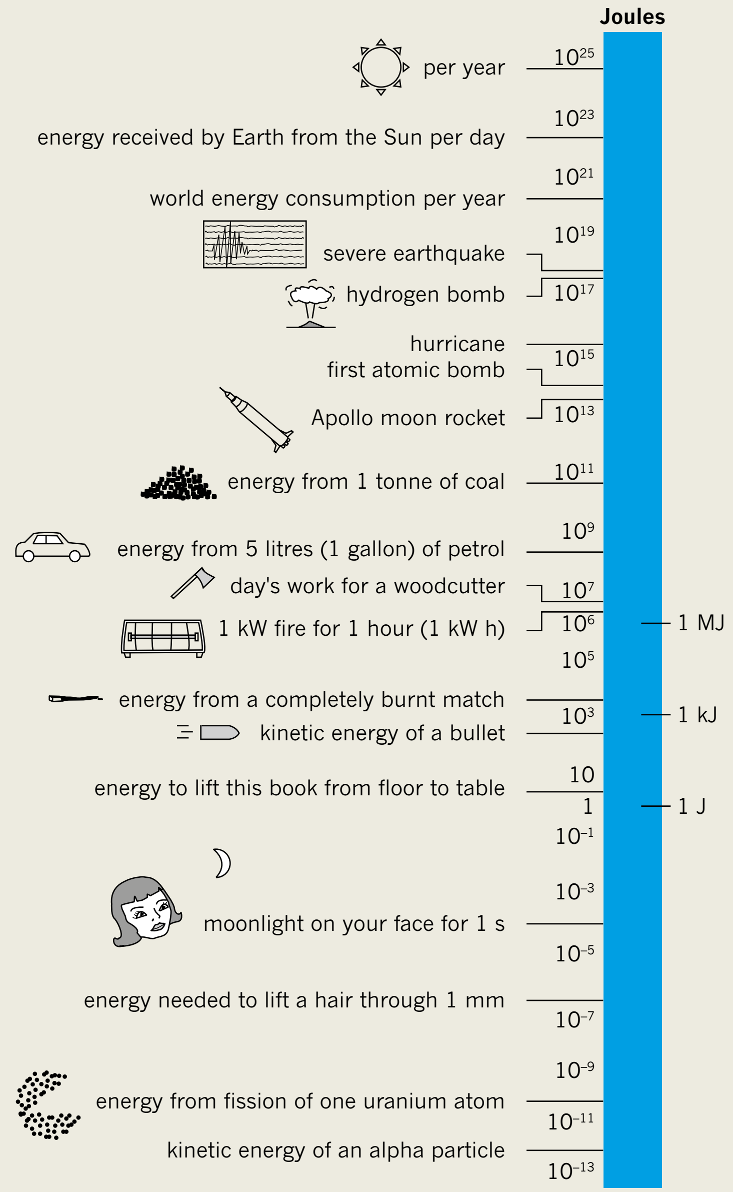


FIGURE 1.1.3 The comparative amounts of energy available from several sources.

- 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.

Using 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 (78%) and oxygen (21%) molecules in the hot air gain energy and so move a lot faster. The air in the balloon becomes less dense than the surrounding air, causing the balloon to float as seen in Figure 1.1.4.

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

For instance, 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 not to change the overall speed of the particles. In this situation, the temperature does not change. This will be discussed in more detail in Section 1.3 'Latent heat'.

MEASURING TEMPERATURE

Only four centuries ago, there were no thermometers and people described heating effects by 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 take into account 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.

Celsius and Fahrenheit scales

Two of the better known arbitrary temperature scales are the Fahrenheit and Celsius scales. Gabriel Fahrenheit of Germany invented the first mercury thermometer in 1714. While Fahrenheit is used in the United States of America to measure temperature, the system used in most of the countries of the world is the Celsius scale.

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 scale

When developing the absolute temperature scale, the triple point of water provides one reliable fixed point. This is a point where the combination of temperature and air pressure allows all three states of water to coexist. For water, the triple point 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. See Figure 1.1.5 for a comparison of the kelvin and Celsius scales.

- The freezing point of water (0°C) is equivalent to 273.15 K (kelvin). This is often approximated to 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.
- To convert a temperature from degrees Celsius to kelvin: add 273.
- To convert a temperature from kelvin to degrees Celsius: subtract 273.

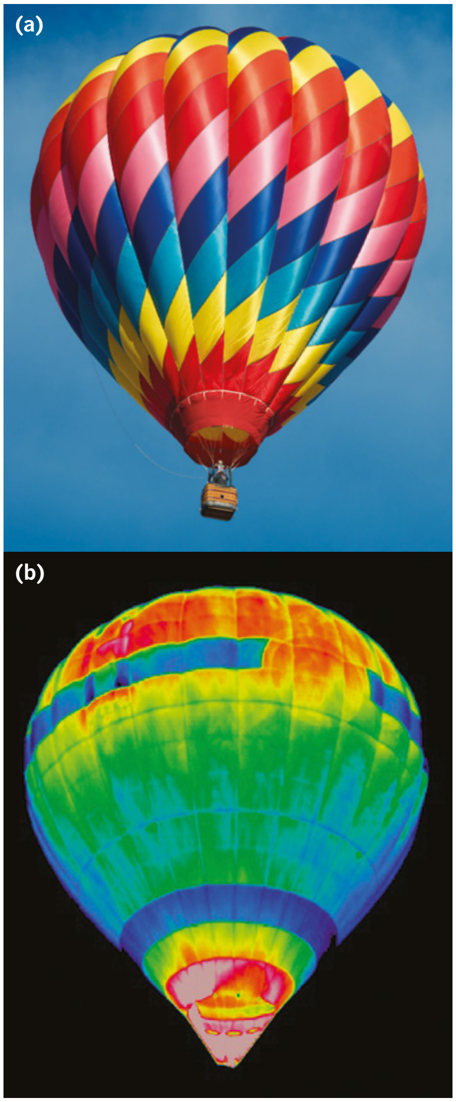


FIGURE 1.1.4 (a) Nitrogen and oxygen molecules gain energy when the air is heated, lowering the density of the air and causing the hot air balloon to rise off the ground. (b) A thermal image shows the temperature of the air inside the balloon.

- 0 degrees Celsius (0°C) is the freezing point of water at standard atmospheric pressure.
- 100°C is the boiling point of water at standard atmospheric pressure.

PHYSICSFILE

Unit conventions in physics

The unit for energy, the joule, is named after James Joule in recognition of his work. When a unit is named after a person, its symbol is usually a capital letter but the unit name is always lower case, e.g. joule (J), newton (N), kelvin (K).

Exceptions are degrees Celsius (°C) and degrees Fahrenheit (°F), which also include a degree symbol.

Units not named after people usually have both the symbol and the name in lowercase, e.g. metre (m), second (s).

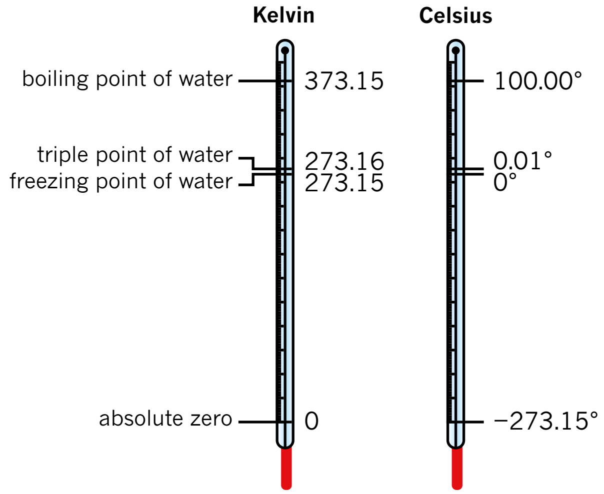


FIGURE 1.1.5 Comparison of the kelvin and Celsius scales. Note that there are no negative values on the kelvin scale.

- Absolute zero = 0 K = −273.15°C
- All molecular motion ceases at absolute zero. This is the coldest temperature possible.

Absolute zero

Experiments indicate that there is a limit to how cold things can get. The kinetic theory suggests that if 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 1.1.6.

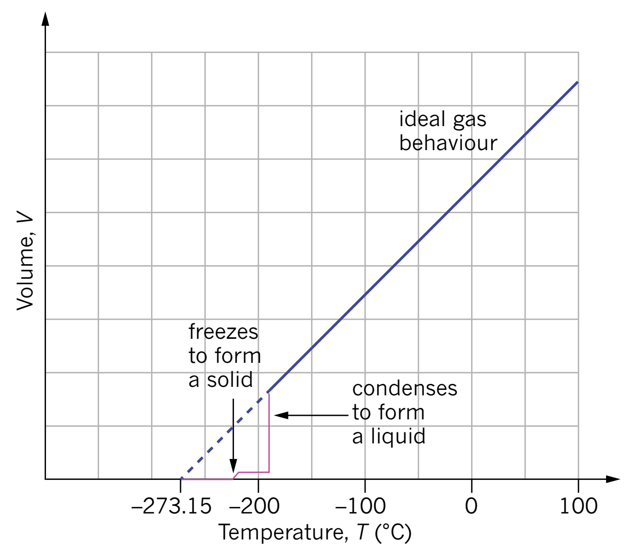


FIGURE 1.1.6 Gases have smaller volumes as they cool. This relationship is linear. Extrapolating (extending) the line to where the volume is zero gives a theoretical value of absolute zero.

PHYSICSFILE

Close to absolute zero

As temperatures get close to absolute zero, atoms start to behave in weird ways. Since the French physicist Guillaume Amontons first proposed the idea of an absolute lowest temperature in 1699, physicists have theorised about the effects of such a temperature and how it could be achieved. The laws of physics dictate that absolute zero itself can be approached but not reached. In 2003, researchers from NASA and MIT, in the United States of America, succeeded in cooling sodium atoms to one billionth of a degree above absolute zero. At this temperature, all elementary particles merge into a single state, losing their separate properties and behaving as a single 'super atom', a state first proposed by Einstein 70 years earlier.

THE LAWS OF THERMODYNAMICS

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

The first, second and third laws had been known and understood for some time. Then another law was also determined. This final law was considered to be 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

The zeroth law of thermodynamics relates to **thermal equilibrium** and **thermal contact** and allows temperature to be defined.

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.

Thermal equilibrium is when two objects in thermal contact stop having a flow of energy between them. If a frozen piece of steak is placed in a container of warm water, energy is transferred from the water to the steak. The steak gains energy and warms up. The water loses energy and cools down. Eventually the transfer of energy between the steak and water will stop. This point is called thermal equilibrium, and the steak and water will be at the same temperature.

> 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.

Two objects in thermal equilibrium with each other must be at the same temperature.

The first law of thermodynamics

The first law of thermodynamics states that energy simply changes 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 ($-Q$), minus the work done on ($-W$) or by ($+W$) 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 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 will 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 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 whether the internal energy increases or decreases will depend on the magnitude of the energy into the system compared to the magnitude of the energy out of the system.

Worked example 1.1.1

CALCULATING THE CHANGE IN INTERNAL ENERGY

A 1 L beaker of water has 25 kJ of work done on it and also loses 30 kJ of thermal energy to the surroundings. What is 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 negative.	$\Delta U = Q - W$ $= -30 - (-25)$
Note that the units are kJ, so express the final answer in kJ.	$\Delta U = -5$ kJ

Worked example: Try yourself 1.1.1

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.

1.1 Review

SUMMARY

- The kinetic particle theory proposes that all matter is made of atoms or molecules (particles) that are in constant motion.
- In solids, the 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.
- 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 Celsius to kelvin: add 273; to convert from kelvin to Celsius: subtract 273.
- 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.
- 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: $\Delta U = Q - W$.

1.1 Review *continued*

KEY QUESTIONS

1 Which of the following is true of a solid?
- **A** Particles are moving around freely.
- **B** Particles are not moving.
- **C** Particles are vibrating in constant motion.
- **D** A solid is not made up of particles.

2 **a** An uncooked chicken is placed into an oven that has been preheated to 180°C. Which of the following statements describe what happens as soon as the chicken is placed in the oven? (More than one answer is possible.)
- **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. 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

3 Which of the following temperature(s) cannot possibly exist? (More than one answer is possible.)
- **A** 1 000 000°C
- **B** −50°C
- **C** −50 K
- **D** −300°C

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.

5 Covert the following temperatures:
- **a** 30°C into kelvin
- **b** 375 K into degrees Celsius.

6 Tank A is filled with hydrogen gas at 0°C and another tank, B, is filled with hydrogen gas at 300 K. Describe the difference in the average kinetic energy of the hydrogen particles in each tank.

7 Sort the following temperatures from coldest to hottest:

freezing point of water
100 K
absolute zero
−180°C
10 K

8 A hot block of iron does 50 kJ of work on a cold floor. The block of iron also transfers 20 kJ of heat energy to the air. Calculate the change in energy (in kJ) of the iron block.

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 J of energy to its surroundings but gains 250 J of energy overall. How much work did the scientist do?

1.2 Specific heat capacity

A small amount 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 get hotter faster than a wooden object. Large heaters warm rooms faster than small ones.

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

CHANGING TEMPERATURE

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 depends on a number of factors.

The greater the mass of a substance, the greater the energy required to change the kinetic energy of all the particles. So, the heat required to raise the temperature by a given amount is proportional to the mass of the substance.

$\Delta Q \propto m$

where ΔQ is the heat energy transferred in joules (J)

m is the mass of material being heated in kilograms (kg).

The more heat that is transferred to a substance, the more the temperature of that substance increases. The amount of energy transferred is therefore proportional to the change in temperature.

$\Delta Q \propto \Delta T$

where ΔT is the change in temperature in °C or K.

Heating experiments using different materials will confirm that these relationships hold true regardless of the material being heated. However, heating the same masses of different materials will show that the amount of energy required to heat a given mass of a material through a particular temperature change also depends on the nature of the material being heated. For example, a volume of water requires more energy to change its temperature by a given amount compared with the same volume of methylated spirits. For some materials, temperature change occurs more easily than for others.

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.

Combining these observations, the amount of energy added to or removed from the substance is proportional to the change in its temperature, its mass and its specific heat capacity (provided a material does not change state). The **specific heat capacity** of a material changes when the material changes state.

As an equation:

$Q = mc\Delta T$

where Q is the heat energy transferred in joules (J)

m is the mass in kilograms (kg)

ΔT is the change in temperature in °C or K

c is the specific heat capacity of the material (J kg^{-1} K^{-1}).

Table 1.2.1 lists the specific heat capacities for some common materials. You can see that it also lists the average value for the human body, which takes into account the various materials within the body and the percentage that each material contributes to the body's total mass.

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

TABLE 1.2.1 Approximate specific heat capacities of common substances.

Worked example 1.2.1

CALCULATIONS USING SPECIFIC HEAT CAPACITY

A hot water tank contains 135 L of water. Initially the water is at 20°C. Calculate the amount of energy that must be transferred to the water to raise the temperature to 70°C.

Thinking	Working
Calculate the mass of water. 1 L of water = 1 kg	Volume = 135 L so mass of water = 135 kg
ΔT = final temperature − initial temperature	$\Delta T = 70 - 20 = 50°C$
From the table of specific heat capacities on page 10, $c_{water} = 4200\ J\ kg^{-1}\ K^{-1}$. Use the equation $Q = mc\Delta T$.	$Q = mc\Delta T$ $= 135 \times 4200 \times 50$ $= 28\,350\,000\ J$ $= 28\ MJ$

Worked example: Try yourself 1.2.1

CALCULATIONS USING SPECIFIC HEAT CAPACITY

A bath contains 75 L of water. Initially the water is at 50°C. Calculate the amount of energy that must be transferred from the water to cool the bath to 30°C.

Worked example 1.2.2

COMPARING SPECIFIC HEAT CAPACITIES

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

Thinking	Working
See Table 1.2.1 for the specific heat capacities of water in different states.	$c_{water} = 4200\ J\ kg^{-1}\ K^{-1}$ $c_{ice} = 2100\ J\ kg^{-1}\ K^{-1}$
Divide the specific heat of water by the specific heat of ice.	Ratio $= \frac{c_{water}}{c_{ice}}$ $= \frac{4200}{2100}$
Note that ratios have no units since the unit of each quantity is the same and cancels out.	Ratio = 2

Worked example: Try yourself 1.2.2

COMPARING SPECIFIC HEAT CAPACITIES

What is the ratio of the specific heat capacity of liquid water to that of steam?

PHYSICSFILE

The mass of water

Since water is a familiar material, many of the examples in this section use it as the liquid being heated. One kilogram of pure water has a volume of 1 litre at 4°C.

PHYSICSFILE

Specific heat capacity of water

One of the notable values in the table of specific heat capacities is the high value for water. It is 10 times, or an order of magnitude, higher than those of most metals listed. 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 in order to determine how the ability of oceans to store large amounts of energy may affect climate change.

1.2 Review

SUMMARY

- 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):
  $Q = mc\Delta T$
  where Q is the heat energy transferred in joules (J)
  m is the mass of material being heated in kilograms (kg)
  ΔT is the change in temperature (°C or K)
  c is the specific heat capacity of the material (J kg^{-1} K^{-1}).
- A substance will have different specific heat capacities at different states (solid, liquid, gas).

KEY QUESTIONS

1 Equal masses of water and aluminium are heated through the same temperature range. Using the values of c from Table 1.2.1 on page 10, which material requires the most energy to achieve this result?

2 Which has the most thermal energy: 10 kg of iron at 20°C or 10 kg of aluminium at 20°C?

3 100 mL of water is heated to change its temperature from 15°C to 20°C. How much energy is transferred to the water to achieve this temperature change?

4 150 mL of water is heated from 10°C to 50°C. What amount of energy is required for this temperature change to occur?

5 For a 1 kg block of aluminium, x J of energy are needed to raise the temperature by 10°C. How much energy, in J, is needed to raise the temperature by 20°C?

6 Equal amounts of energy are absorbed by equal masses of aluminium and water. What is the ratio of the temperature rise of the aluminium to that of water?

7 Which one or more of the following statements about specific heat capacity is true?

A All materials have the same specific heat capacity when in solid form.

B The specific heat capacity of a liquid form of a material is different from that of the solid and gas forms.

C Good conductors of heat generally have high specific heat capacities.

D Specific heat capacity is independent of temperature.

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

9 A cup holds 250 mL of water at 20°C. 10.5 kJ of heat energy is transferred to the water. What temperature does the water reach after the heat is transferred?

10 A block of iron is left to cool. After cooling for a short time, 13.2 kJ of energy has been transferred away from the block of iron and its temperature has decreased by 30°C. What is the mass of the block of iron?

1.3 Latent heat

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 changes state, its temperature remains constant. The energy used in, say, melting ice 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 heating curve for water shown in Figure 1.3.1. This graph shows how the temperature of water changes as energy is added at a constant rate. Although the rate at which the energy is added is constant, the increase in temperature is not always constant. There are sections of increasing temperature, and sections where the temperature remains unchanged (the horizontal sections) while the material changes state. Temperature remains constant during the change in state from ice to liquid water and again from liquid water to steam.

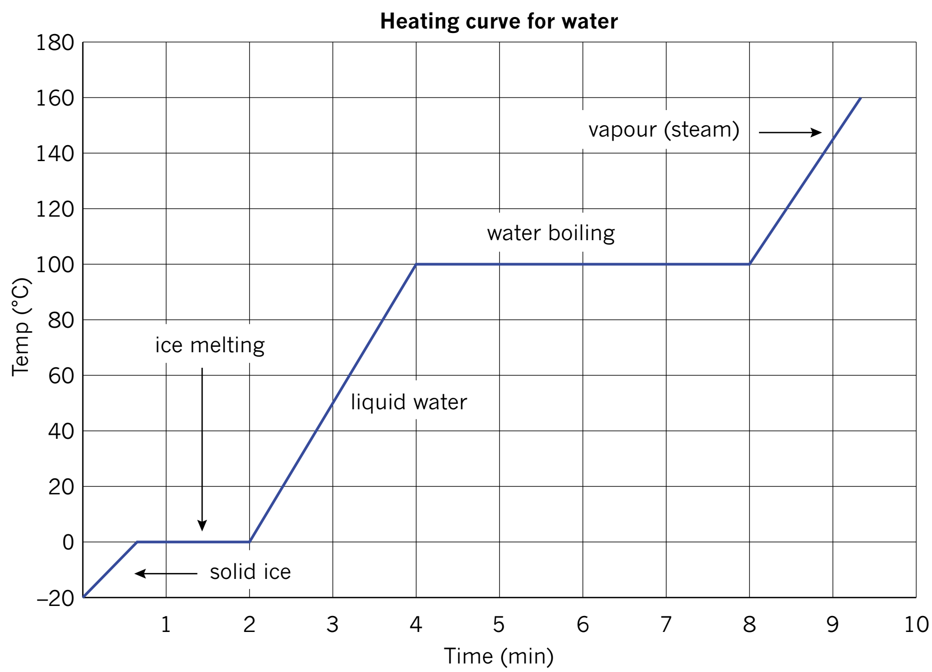


FIGURE 1.3.1 A heating curve for water.

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 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)

As energy is transferred to a solid, the temperature of the solid increases. 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 as all the extra energy supplied is used in reducing these forces between particles.

The amount of energy required to melt a solid is exactly 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 particular solid.

For a given mass of a substance:

heat energy transferred = mass of substance × specific latent heat of fusion

$Q = mL_{fusion}$

where Q is the heat energy transferred in joules (J)

m is the mass in kilograms (kg)

L_{fusion} is the latent heat of fusion in J kg^{-1}.

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 shown in Table 1.3.1.

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

TABLE 1.3.1 The latent heats of fusion for some common materials.

Worked example 1.3.1

LATENT HEAT OF FUSION

How much energy must be removed from 2.5 L of water at 0°C to produce a block of ice at 0°C? Express your answer in kJ.

Thinking	Working
Cooling from liquid to solid involves the latent heat of fusion, where the energy is removed from the water. Calculate the mass of water involved.	1 L of water = 1 kg, so 2.5 L = 2.5 kg
Use Table 1.3.1 to find the latent heat of fusion for water.	$L_{fusion} = 3.34 \times 10^5$ J kg^{-1}
Use the equation $Q = mL_{fusion}$.	$Q = mL_{fusion}$ $= 2.5 \times 3.34 \times 10^5$ $= 8.35 \times 10^5$ J
Convert to kJ.	$Q = 8.35 \times 10^2$ kJ

Worked example: Try yourself 1.3.1

LATENT HEAT OF FUSION

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)

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 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 exactly 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 particular substance.

For a given mass of a substance:

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

$Q = mL_{vapour}$

where Q is the heat energy transferred in joules (J)

m is the mass in kilograms (kg)

L_{vapour} is the latent heat of vaporisation (J kg^{-1}).

Note that, in just about every case, the latent heat of vaporisation of a substance will be different to the latent heat of fusion for that substance. Some latent heat of vaporisation values are listed in Table 1.3.2.

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

Substance	Melting point (°C)	L_{vapour} (J kg^{-1})
water	100	22.5×10^5
oxygen	−183	2.2×10^5
lead	1750	9.0×10^5
ethanol	78	8.7×10^5
silver	2193	23.0×10^5

TABLE 1.3.2 The latent heat of vaporisation of some common materials.

Worked example 1.3.2

CHANGE IN TEMPERATURE AND STATE

50 mL of water is heated from a room temperature of 20°C to its boiling point at 100°C. It is boiled at this temperature until it is completely evaporated. How much energy in total was required to raise the temperature and boil the water?

Thinking	Working
Calculate the mass of water involved.	50 mL of water = 0.05 kg
Find the specific heat capacity of water (see Table 1.2.1).	$c = 4200$ 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.	$Q = mc\Delta T$ $= 0.05 \times 4200 \times (100 - 20)$ $= 16\,800$ J
Find the specific latent heat of vaporisation of water.	$L_{vapour} = 22.5 \times 10^5$ J kg^{-1}
Use the equation $Q = mL_{vapour}$ to calculate the latent heat required to boil water.	$Q = mL_{vapour}$ $= 0.05 \times 22.5 \times 10^5$ $= 112\,500$ J
Find the total energy required to raise the temperature and change the state of the water.	Total $Q = 16\,800 + 112\,500$ $= 129\,300$ J (or 1.29×10^5 J)

PHYSICSFILE

Extinguishing fire

The latent heat of vaporisation of water is very high. This is due to the molecular structure of the water. This characteristic of water makes it very useful for extinguishing fires. That's because water can absorb vast amounts of thermal energy before it evaporates. By pouring water onto a fire, energy is transferred away from the fire to heat the water. Then, even more (in fact much more) heat is transferred away from the fire to convert the water into steam.

Worked example: Try yourself 1.3.2

CHANGE IN TEMPERATURE AND STATE

3 L of water is heated from a fridge temperature of 4°C to its boiling point at 100°C. It is boiled at this temperature until it is completely evaporated. 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, are able to 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 lid off bottles of these liquids. They are often stored in narrow-necked bottles for this reason.

The rate of evaporation of a liquid depends 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 liquid's surface, evaporation is more rapid.

Whenever evaporation occurs, higher-energy particles escape the surface of the liquid, leaving the lower-energy particles behind, as is shown in Figure 1.3.2. As a result, the average kinetic energy of the particles remaining in the liquid drops and the temperature decreases. Humans use this cooling principle when sweating to stay cool. When rubbing alcohol is dabbed on your arm before an injection, the cooling of the volatile liquid numbs your skin.

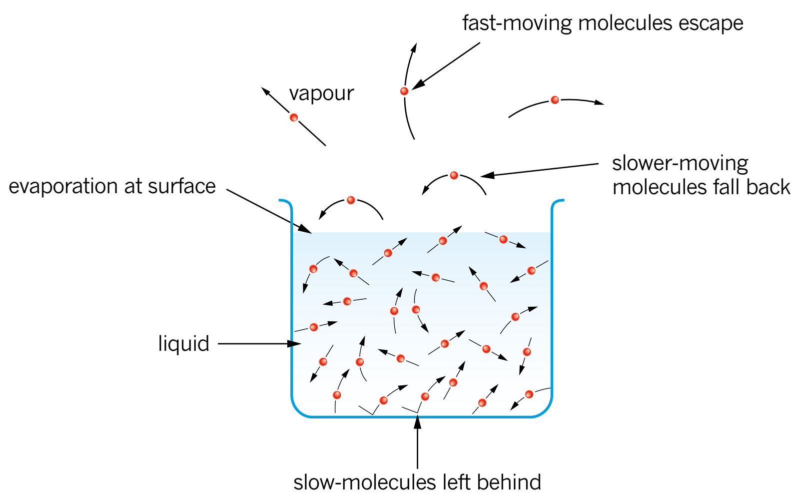


FIGURE 1.3.2 Fast-moving molecules with high kinetic energy can escape the liquid, leaving molecules with lower kinetic energy behind.

Other applications of evaporation such as evaporative coolers will be discussed in more detail in Chapter 2.

1.3 Review

SUMMARY

- When a solid material changes state, energy is needed to separate the particles by overcoming the attractive forces between the particles.
- 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
  $Q = mL$
  where Q is the energy transferred in joules (J)
  m is the mass in kilograms (kg)
  L is the latent heat (J kg^{-1}).
- The latent heat of fusion, L_{fusion}, is the energy required to change 1 kg of a material between the solid and liquid states.
- The latent heat of vaporisation, L_{vapour}, is the energy required to change 1 kg of a material between the liquid and gaseous states.
- The latent heat of fusion of a material will be different to (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 a breeze.

KEY QUESTIONS

Refer to the values in Table 1.3.1 and Table 1.3.2 on pages 14 and 15. You may also need to refer to Table 1.2.1 on page 10.

The following information relates to questions 1 to 5.

Figure 1.3.3 represents the heating curve for mercury, a metal that is a liquid at normal room temperature. Thermal energy is added to 10 g of solid mercury, initially at a temperature of –39°C, until all of the mercury has evaporated.

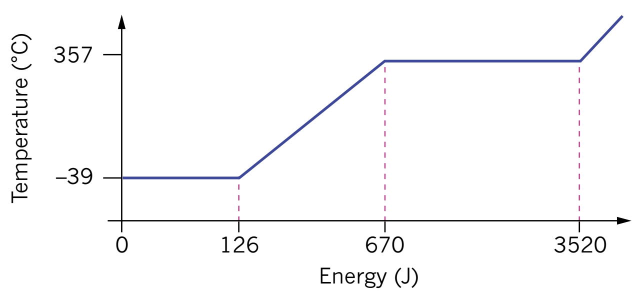


FIGURE 1.3.3 A heating curve of mercury.

1. Why does the temperature remain constant during the first part of the graph?
2. What is the melting point of mercury, in degrees Celsius?
3. What is the boiling point of mercury, in degrees Celsius?
4. From the graph, what is the latent heat of fusion of mercury?
5. From the graph, what is the latent heat of vaporisation of mercury?
6. How much heat energy must be transferred away from 100 g of steam at 100°C to change it completely to a liquid?
7. How many kJ of energy are required to melt exactly 100 g of ice initially at −4.00°C? Assume no loss of energy to surroundings.
8. Which of the following explains why hot water in a spa-pool evaporates more rapidly than cold water?
   A Hot water molecules have less energy than cold water molecules.
   B Hot water molecules have more energy than cold water molecules.
   C Hot water forms water vapour and bubbles to the surface.
   D Hot water dissolves into the pool material more rapidly.
9. A painter spills some mineral turpentine onto a concrete floor. After a minute, most of the liquid is gone. Which of the following is correct?
   A Most of the liquid has boiled, becoming hotter.
   B Most of the liquid has evaporated and the remaining liquid becomes warmer as it does so.
   C Most of the liquid has evaporated with no change in temperature of the remaining liquid.
   D Most of the liquid has evaporated and the remaining liquid becomes colder as it does so.

1.4 Conduction

FIGURE 1.4.1 Emperor penguin chicks avoid heat loss through conduction by sitting on the adult's feet. In this way they avoid contact with the ice.

If two objects are at different temperatures and 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 1.4.1 shows how, by preventing the chick's thermal contact with the cold ice, this adult is able to protect the vulnerable penguin offspring.

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

- conduction
- convection
- radiation.

This section focuses on conduction.

CONDUCTORS AND INSULATORS

Conduction is the process by which heat is transferred from one place to another without the net movement of particles (atoms or molecules). 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; or if a person holds an ice cube, then heat will travel from their hand to the ice.

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

Materials that conduct heat readily are referred to as good **conductors**. Materials that are poor conductors of heat are referred to as **insulators**. An example of a good conductor and a good insulator can be seen in Figure 1.4.2.

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.

Conduction can happen in two ways:

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

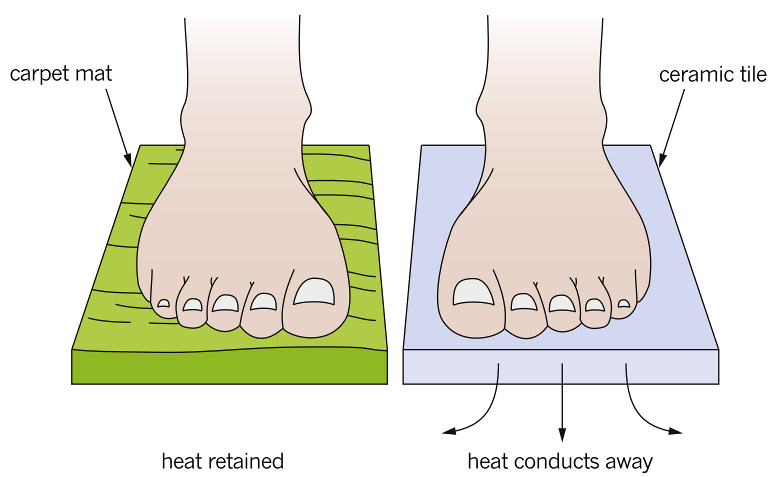


FIGURE 1.4.2 Ceramic floor tiles are good conductors of heat. They conduct heat away from the foot readily and so your feet feel cold on tiles. The carpet mat is a thermal insulator. Thermal energy from the foot is not transferred away as quickly and so your foot doesn't feel as cold.

THERMAL TRANSFER BY COLLISION

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 (see Figure 1.4.3).

The process can be quite slow since the mass of the particles is relatively large and the vibrational velocities are fairly low. Materials for which this method of conduction is the only means of heat transfer are likely to be poor conductors of heat or even thermal insulators. Materials such as glass, wood and paper are poor conductors of heat.

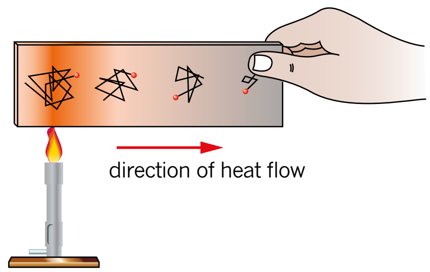


FIGURE 1.4.3 Thermal energy is passed on by collisions between adjacent particles.

THERMAL TRANSFER BY FREE ELECTRONS

Some materials, particularly metals, have electrons that are not directly involved in any one particular chemical bond. Therefore, these electrons are free to move throughout the lattice of positive ions.

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 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 1.4.1 highlights the difference in conductivity in metals compared with other substances.

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

TABLE 1.4.1 Thermal conductivities of some common materials.