UNIT 1

AREA OF STUDY 1

CHAPTER 1 Heat, temperature and internal energy
CHAPTER 2 Energy in transit
CHAPTER 3 The physics of climate change

AREA OF STUDY 2

CHAPTER 4 Current electricity
CHAPTER 5 Circuit analysis
CHAPTER 6 Using electricity

AREA OF STUDY 3

CHAPTER 7 Radioactivity and the nucleus
CHAPTER 8 Subatomic particles
CHAPTER 9 The origin of atoms
Incoming radiation is reflected off ice back into space, but is absorbed by the water. How will increasing air and ocean temperatures change this scene?

**REMEMBER**
Before beginning this chapter you should be able to:
- recall the different forms of energy
- recognise changes between different forms of energy
- use a thermometer
- describe the particle model of matter.

**KEY IDEAS**
After completing this chapter you should be able to:
- convert temperature between degrees Celsius and Kelvin
- describe the kinetic particle model of matter
- explain internal energy as the energy associated with random disordered motion of molecules
- describe temperature with reference to the average translational kinetic energy of the atoms and molecules
- describe the Zeroth Law of Thermodynamics as two bodies in contact with each other coming to a thermal equilibrium
- explain the First Law of Thermodynamics, \( \Delta U = Q - W \), and apply it to simple situations
- describe the specific heat capacity of a substance and use it to calculate the energy required to raise its temperature with \( Q = mc \Delta T \)
- describe the latent heat of a substance and use it to calculate the energy required to change its state with \( Q = mL \)
- explain why cooling results from evaporation using the kinetic energy model of matter.
Measuring temperature

Our bodies tell us when it is hot or cold. Our fingers warn us when we touch a hot object. However, for all that, our senses are not reliable.

Try this at home: Place three bowls of water in front of you. Put iced water in the bowl on the left, water hot enough for a bath in the bowl on the right, and room temperature water in the one in the middle. Place a hand in each of the two outer bowls, leave them there for a few minutes, then place both hands in the middle bowl. As you would expect, your left hand tells you the water is warmer, while your right hand tells you it is colder.

Thermometers were designed as a way to measure temperature accurately. A good thermometer needs a material that changes in a measurable way as its temperature changes. Many materials, including water, expand when heated, so the first thermometer, built in 1630, used water in a narrow tube with a filled bulb at the bottom. The water rose up the tube as the bulb was warmed.

German physicist Daniel Fahrenheit replaced the water with mercury in 1724. Liquid thermometers now use alcohol with a dye added. Fahrenheit developed a scale to measure the temperature, using the lowest temperature he could reach, an ice and salt mixture as zero degrees, and the temperature of the human body as 100 degrees Fahrenheit. He also showed that a particular liquid will always boil at the same temperature. Swedish astronomer Anders Celsius developed another temperature scale in 1742, which is the one we use today. Celsius used melting ice and steam from boiling water to define 0 °C and 100 °C for his scale.

A third temperature scale was proposed in 1848 by William Thomson, later to be ennobled as Lord Kelvin. He proposed the scale based on the better understanding of heat and temperature that had developed by that time (see page 10). This scale uses the symbol ‘K’ to stand for ‘Kelvin’.

Table 1.1: Some temperatures on the Kelvin and Celsius scales

<table>
<thead>
<tr>
<th>Event</th>
<th>Temperature</th>
</tr>
</thead>
<tbody>
<tr>
<td>Absolute zero</td>
<td>0</td>
</tr>
<tr>
<td>Helium gas liquefies</td>
<td>4</td>
</tr>
<tr>
<td>Lead becomes a superconductor</td>
<td>7</td>
</tr>
<tr>
<td>Nitrogen gas liquefies</td>
<td>63</td>
</tr>
<tr>
<td>Lowest recorded air temperature on the Earth’s surface (Vostok, Antarctica)</td>
<td>184</td>
</tr>
<tr>
<td>Mercury freezes</td>
<td>234</td>
</tr>
<tr>
<td>Water freezes</td>
<td>273</td>
</tr>
<tr>
<td>Normal human body temperature</td>
<td>310</td>
</tr>
<tr>
<td>Highest recorded air temperature on the Earth’s surface (Death Valley, USA)</td>
<td>330</td>
</tr>
<tr>
<td>Mercury boils</td>
<td>630</td>
</tr>
<tr>
<td>Iron melts</td>
<td>1535</td>
</tr>
<tr>
<td>Surface of the Sun</td>
<td>5778</td>
</tr>
</tbody>
</table>

Other materials, including gases and metals, also expand with temperature and are used as thermometers. A bimetallic strip is two lengths of different metals, usually steel and copper, joined together. The two metals expand at
different rates, so the strip bends one way as the temperature rises, or the other as it cools. A bimetallic strip can be used as a thermometer, a thermostat or as a compensating mechanism in clocks.

Other properties that change with temperature that can be employed in designing a thermometer are:

- electrical resistance of metals, which increases with temperature
- electrical voltage from a thermocouple, which is two lengths of different metals with their ends joined; if one end is heated, a voltage is produced
- colour change; liquid crystals change colour with temperature
- colour emitted by a hot object; in steel making, the temperature of hot steel is measured by its colour.

This steel is nearly 1000 °C and has recently been poured in a mould to shape it. The steel will continue to glow until it has cooled to about 400 °C.

This thermocouple is connected to a voltmeter which reads differing voltages as the thermocouple changes temperature.

This liquid crystal thermometer indicates body temperature when the liquid crystals change colour. The thermometer is registering 37 °C.
What is the difference between heat and temperature?

If you mix a beaker of cold water at 10 °C with a beaker of hot water at 50 °C, you expect the final temperature to be about 30 °C. But if you mix the beaker of cold water with a jug of hot water, the final temperature will be a lot closer to 50 °C.

Part A:

Part B:

Obviously if an object has more mass, it contains more heat, but how are temperature and heat related?

Consider these two analogies to temperature and heat:

(a) If you drop a marble on your big toe from a height of 1 metre, you would notice it, but it would not hurt. However, if you dropped a 1 kg mass from only 10 cm it would hurt.

   The height is to temperature as the impact is to heat.

(b) If you rub your shoes across a carpet, you can generate a voltage as high as 10,000 volts, but the electric shock is no more than a twitch. However, a 6-volt car battery can deliver a charge to start the engine.

   The voltage is to temperature as the spark is to heat.
These analogies don’t really help to explain what heat is. If an object cools down, there seems to be no physical difference other than the drop in temperature. The object does not weigh less because it has lost heat! What is being transferred when a hot object warms up a cold object?

In 1798 Benjamin Thompson, later to be called Count Rumford, conducted an experiment on the nature of heat. The barrel of a cannon is made by drilling a cylindrical hole in a solid piece of metal. Rumford observed that metal and the drill became quite hot. He devised an experiment to investigate the source of the heat and how much heat is produced. Rumford put the drill and the end of the cannon in a wooden box filled with water. He measured the mass of water and the rate at which the temperature rose. He showed that the amount of heat produced was not related to the amount of metal that was drilled out. He concluded that the amount of heat produced depended only on the work done against friction. He said that heat was in fact a form of energy, not an invisible substance that is transferred from hot objects to cold objects. Instead a hot object had heat energy, in the same way as a moving object has kinetic energy or an object high off the ground has gravitational potential energy.

**Kinetic energy** is the energy associated with the movement of objects. Like all forms of energy, kinetic energy is a scalar quantity.

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Count Rumford was born Benjamin Thompson in Massachusetts in 1753. By the age of 16 he was conducting experiments on heat. For 1775, when the American War of Independence began, he was already a wealthy man and of some standing in his community. He joined the British side of the war, becoming a senior advisor. While with the army, he also investigated and published a paper on the force of gunpowder.

At the end of the war, he moved to England, where he was known as a research scientist. A few years later he moved to Bavaria, in what is now southern Germany, and spent 11 years there. He moved in royal circles and eventually became Bavaria’s Army Minister, tasked with reorganising the army. As part of those duties, he investigated methods of cooking, heating and lighting. He developed a soup, now called Rumford’s soup, as a nutritious ration for soldiers. He also used the soup to establish soup kitchens for the poor throughout Bavaria. For his services he was made a Count of the Holy Roman Empire, taking the name ‘Rumford’ from his birthplace.

On return to England, his activities included: (i) redesigning an industrial furnace, which revolutionised the production of quicklime, a component of cement, and also used for lighting (‘limelight’); (ii) redesigning the domestic fireplace to narrow the chimney at the hearth to increase the updraught, resulting in greater efficiency and no smoke coming back into the room; and (iii) inventing thermal underwear, a kitchen range and a drip coffeepot.

With Joseph Banks and others, Rumford established the Royal Institution (RI) in London as a scientific research establishment with a strong emphasis on public education. Initial funding came from the ‘Society for Bettering the Conditions and Improving the Comforts of the Poor’, with which Count Rumford was centrally involved. Famous scientists in its early years included Humphrey Davy and Michael Faraday. Fifteen Nobel Prize winners have worked at the RI, and 10 chemical elements were discovered there.

**Revision question 1.1**

Devise an experiment to investigate the heat generated when two hands are rubbed together.
Rumford’s ideas about heat were not taken up for a few decades. But in 1840 James Prescott Joule conducted a series of experiments to find a quantitative link between mechanical energy and heat. In other words, how much energy is required to increase the temperature of a mass by 1 °C?

Joule used different methods and compared the results:

- Using gravity: A falling mass spins a paddle wheel in an insulated barrel of water, raising the temperature of the water.
- Using electricity: Mechanical work is done turning a dynamo to produce an electric current in a wire, which heats the water.
- Compressing a gas: Mechanical work is used to compress a gas, which raises the gas’s temperature.
- Using a battery: Chemical reactions at the battery terminals produce a current, which heats the water.
- Using gravity: Measure the temperature of water at the top and bottom of a waterfall.

Joule obtained approximately identical answers for all methods. This confirmed heat as a form of energy. To honour his achievement, the SI unit of energy is the joule (J). The unit, joule, is used to measure the kinetic energy of a runner, the light energy in a beam, the chemical energy stored in a battery, the electrical energy in a circuit, the potential energy in a lift on the top floor and the heat energy when water boils.

One joule is approximately the amount of energy needed to lift a 100 g apple through a height of 1 m. The usual metric prefixes make the use of the unit more convenient. For example:

- 1 kJ (kilojoule) = 10^3 J
- 1 MJ (megajoule) = 10^6 J
- 1 GJ (gigajoule) = 10^9 J

The chemical energy available from a bowl of breakfast cereal is usually hundreds of thousands of joules and is more likely to be listed on the packet in kilojoules. The amount of energy needed to boil an average kettle full of cold water is about 500 kJ.

Examples of 1 joule include:
- the kinetic energy of a tennis ball moving at about 6 m/s
- the heat energy needed to raise the temperature of 1 gram of dry air by 1 °C
- the heat energy needed to raise the temperature of 1 gram of water by 0.24 °C
- the energy released when an apple falls 1 m to the ground
- the amount of sunlight hitting a square centimetre every 10 seconds when the Sun is directly above
- the amount of sound energy entering your eardrum at a loud concert over 3 hours
- the amount of electrical energy used by a plasma TV screen while on standby every 2.5 seconds
- the energy released by the combustion of 18 micrograms of methane.

Explaining heat: the kinetic theory of matter

The kinetic theory of matter, which considers all objects as assemblies of particles in motion, is an old one, first described by Lucretius in 55 AD. The kinetic view of matter was developed over time by Hooke, Bernoulli, Boltzmann and Maxwell.

The evidence for the existence of particles includes:
- gases and liquids diffuse, that is, a combination of two gases or two liquids quickly becomes a mixture, for example a dye spreading in water. Even solids can diffuse, if a sheet of lead is clamped to a sheet of gold, over time the metals merge to a depth of a few millimetres.
- the mixing of two liquids gives a final volume that is less than the sum of their original volumes
- a solid dissolves in a liquid.
The kinetic theory of matter assumes that:
• all matter is made up of particles in constant, random and rapid motion
• there is space between the particles.

The energy associated with the motion of the particles in an object is called the internal energy of the object. The particles can move and interact in many ways, so there are a number of contributions to the internal energy. For example:

Gases: In a gas made up of single atoms, such as helium, the atoms move around, randomly colliding with each other and the walls of the container. So each atom has some translational kinetic energy.

However, if the gas is made up of molecules with two or more atoms, the molecules can also stretch, contract and spin, so these molecules also have other types of kinetic energy called vibrational and rotational kinetic energy.
Liquids: Like a gas, molecules in a liquid are free to move, but within the confines of the surface of the liquid. There is some attraction between molecules, which means there is some energy stored as molecules approach each other. Stored energy is called potential energy. It is the energy that must be overcome for a liquid to evaporate or boil.

Solids: In a solid, atoms jiggled rather than move around. They have kinetic energy, but they also have a lot of potential energy stored in a strong attractive force that holds the atoms together. This means that a lot of energy is required to melt a solid.

**TABLE 1.2**

<table>
<thead>
<tr>
<th>Internal energy</th>
<th>Movement that is NOT related to temperature</th>
<th>Movement that is related to temperature</th>
</tr>
</thead>
<tbody>
<tr>
<td>Atoms in a gas</td>
<td>None</td>
<td>Moving and colliding</td>
</tr>
<tr>
<td>Molecules in a gas</td>
<td>Spinning, stretching, compressing and bending</td>
<td>Moving and colliding</td>
</tr>
<tr>
<td>Molecules in a liquid</td>
<td>Spinning, stretching, compressing and bending</td>
<td>Moving and colliding</td>
</tr>
<tr>
<td>Atoms in a solid</td>
<td>Pulling and pushing</td>
<td>Jiggling</td>
</tr>
<tr>
<td>Energy types</td>
<td>Other types of kinetic energy, potential energy</td>
<td>Translational kinetic energy</td>
</tr>
</tbody>
</table>

**Temperature** is a measure of the average translational kinetic energy of particles. The other contributions to the internal energy do not affect the temperature. This becomes important when materials melt or boil because the added heat must go somewhere, but the temperature does not change.

The kinetic theory of matter is the origin of the Kelvin temperature scale. If temperature depends on the movement of particles, then the slower they...
move, the lower the temperature. When the particles stop moving, the temperature will be the lowest that is physically possible. This temperature was adopted as absolute zero. But how do we measure it and what is its value?

In the early 1800s gases were a good material to work with to explore the nature of matter. An amount of gas in a glass vessel could be heated and the variables of temperature, volume and pressure to keep the volume fixed could be easily measured. Joseph Gay-Lussac and Jacques Charles independently investigated how the volume of gases changed with temperature if they were kept under a constant pressure. They found that all gases kept at constant pressure expand or contract by $\frac{1}{273}$ of their volume at $0 \degree C$ for each Celsius degree rise or fall in temperature.

From that result you can conclude that if you cooled the gas, and it stayed as a gas and did not liquefy, you could cool it to a low enough temperature that its volume reduced to zero. The temperature would be absolute zero. According to their experiments, absolute zero was $-273 \degree C$. Nowadays more accurate experiments put the value at $-273.15 \degree C$.

In degrees Kelvin, absolute zero is 0 K. The increments in the Kelvin temperature scale are the same size as those in the Celsius scale, so if the temperature increased by 5 \degree C, it also increased by 5 K. The conversion formula between the two temperature scales is:

$$\text{degrees Kelvin} = \text{degrees Celsius} + 273$$

**Sample problem 1.1**

What is the Kelvin temperature at which ice melts?

*Solution:* Ice melts at 0 \degree C, so the equivalent Kelvin temperature is $0 + 273 = 273$ K.

*Note:* In 1968, the international General Conference on Weight and Measures decided that Kelvin temperatures do not use the \degree symbol as do Celsius and Fahrenheit temperatures.

**Revision question 1.2**

(a) Carbon dioxide sublimes, that is, goes directly from solid to gas, at $-78.5 \degree C$. What is this temperature in degrees Kelvin?

(b) The temperature of the surface of the Mars was measured by the Viking lander and ranged from 256 K to 166 K. What are the equivalent temperatures in degrees Celsius?
**Thermal equilibrium**

Energy is always transferred from a region of high temperature to a region of lower temperature until both regions reach the same temperature. When the temperature is uniform, a state of thermal equilibrium is said to exist.

So when a hot nail is dropped into a beaker of cold water, energy will be transferred from the hot nail into the water even though the hot nail has less total internal energy than the water. When thermal equilibrium is reached, the temperature of both the water and the nail is the same. The particles in the water and the particles in the nail have the same amount of random translational kinetic energy. The figure below shows how the kinetic particle model can be used to explain the direction of energy transfer in the beaker.

When you swim in a cold pool, energy is transferred from your body into the water. The water has much more total internal energy than your body because there is so much of it. However, the particles in your body have more random translational kinetic energy that can be transferred to the particles of cold water. Hopefully, you would not remain in the water long enough for thermal equilibrium to be reached.

What is implicit in the above discussion on thermal equilibrium and internal energy, is the subtle, but important, point made by James Clerk Maxwell that ‘All heat is of the same kind’.
Laws of thermodynamics

Three laws of thermodynamics were progressively developed during the 19th century, but in the 20th century it became apparent that the principle of thermal equilibrium could be seen as the logical underpinning of these three laws. Consequently the Zeroth Law of Thermodynamics became accepted.

Zeroth Law of Thermodynamics

Consider three objects: A, B and C. It is the case that A is in thermal equilibrium with B, and C is also in thermal equilibrium with B. Since ‘All heat is of the same kind’, it follows that A is in thermal equilibrium with C.

In practice this means that all three objects, A, B and C, are at the same temperature, and the law enables the comparison of temperatures.

First Law of Thermodynamics

The First Law of Thermodynamics states that energy is conserved and cannot be created or destroyed. If there is an energy change in a system, all the energy must be accounted for. From a thermodynamics point of view, the internal energy of a substance and any change in it are a crucial part of this accounting exercise.

Consider a volume of air inside a balloon that is placed in direct sunlight. The air inside the balloon will get hotter and the balloon will expand slightly.

The First Law of Thermodynamics says:

\[ \Delta U = Q - W \]

where

- \( \Delta U \) is the change in internal energy in joules
- \( Q \) is the heat energy in joules
- \( W \) is the work done in joules

Using symbols:

\[ \Delta U = Q - W \]

The energy from the Sun heats the air inside the balloon, increasing the kinetic energy of the air molecules. The air molecules lose some of this energy as they repeatedly collide with the wall of the balloon, forcing it outwards.

The First Law of Thermodynamics applies to many situations: cylinders in a car engine, hot air balloons, food consumption, pumping up a tyre and the weather. Consequently, the word ‘system’ is often used as a generic name when discussing thermodynamics.

Note: The words in bold, ‘of’, ‘to’ and ‘by’, and the minus sign are important in the equation as \( Q \) and \( W \) can be either positive or negative.

Guide:

If a system absorbs heat, e.g. energy from sunlight, then \( Q > 0 \).
If a system releases heat, e.g. when you sweat, then \( Q < 0 \).
If a system does work on the surroundings, e.g. hot balloon expands, then \( W > 0 \).
If the surroundings do work on the system, e.g. pumping up a tyre, then \( W < 0 \).

Sample problem 1.2

(a) A balloon is placed in direct sunlight. The sunlight supplies 200 joules of energy to the balloon. The air inside pushes out the balloon surface, doing 50 joules of work. By how much does the internal energy of the air inside change?

(b) While doing some heavy lifting, you do 2500 joules of work on the weights, while releasing 3000 joules of heat. By how much did your internal energy change?
### Solution:

(a) $Q = 200 \text{ J}, W = 50 \text{ J}$, so $\Delta U = 200 - 50 = 150 \text{ J}$. The internal energy of the air in the balloon increased by 150 J.

(b) $Q = -3000 \text{ J}, W = 2500 \text{ J}$, so $\Delta U = -3000 - 2500 = -5500 \text{ J}$. Your internal energy decreased by 5500 J.

### Revision question 1.3

A block of ice is melted by 100 joules of energy. What is the size and the sign of $W$ and $\Delta U$?

### Specific heat capacity

Once the temperature of materials could be accurately measured, it became apparent that, when heated, some materials increased in temperature more quickly than others. The property of the material that describes this phenomenon is called the specific heat capacity and is defined as the amount of energy required to increase the temperature of 1 kg of the substance by 1 °C (or K).

It takes more energy to increase the temperature of water by 1 °C than any other common substance. Water also needs to lose more energy to decrease in temperature. In simple terms, this means that water maintains its temperature well, cooling down and heating up more slowly than other materials.

### Table 1.3 Specific heat capacity of some common substances

<table>
<thead>
<tr>
<th>Substance</th>
<th>Specific heat capacity (J kg$^{-1}$ K$^{-1}$)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Helium</td>
<td>5193</td>
</tr>
<tr>
<td>Water</td>
<td>4200</td>
</tr>
<tr>
<td>Human body (average)</td>
<td>3500</td>
</tr>
<tr>
<td>Cooking oil</td>
<td>2800</td>
</tr>
<tr>
<td>Ethylene glycol (used in car ‘antifreeze’)</td>
<td>2400</td>
</tr>
<tr>
<td>Ice</td>
<td>2100</td>
</tr>
<tr>
<td>Steam</td>
<td>2000</td>
</tr>
<tr>
<td>Fertile topsoil</td>
<td>1800</td>
</tr>
<tr>
<td>Neon</td>
<td>1030</td>
</tr>
<tr>
<td>Air</td>
<td>1003</td>
</tr>
<tr>
<td>Aluminium</td>
<td>897</td>
</tr>
<tr>
<td>Carbon dioxide</td>
<td>839</td>
</tr>
<tr>
<td>Desert sand</td>
<td>820</td>
</tr>
<tr>
<td>Glass (standard)</td>
<td>670</td>
</tr>
<tr>
<td>Argon</td>
<td>520</td>
</tr>
<tr>
<td>Iron and steel (average)</td>
<td>450</td>
</tr>
<tr>
<td>Zinc</td>
<td>387</td>
</tr>
<tr>
<td>Copper</td>
<td>385</td>
</tr>
<tr>
<td>Lead</td>
<td>129</td>
</tr>
</tbody>
</table>

Specific heat capacities differ because of two factors:
- the different contributions to the internal energy by the forms of energy other than translational kinetic energy, and
- the varying mass of individual atoms and molecules.
The internal energy of single-atom gases, such as helium, neon and argon, consists of only translational kinetic energy. So the specific heat capacities should be the same if you account for their difference in mass. Look up the atomic weight for each gas and multiply it by each gas's specific heat capacity and compare your answers.

The quantity of energy, \( Q \), transferred to or from a substance in order to change its temperature is directly proportional to three factors:

- the mass of the substance \((m)\)
- the change in temperature \((\Delta T)\)
- the specific heat capacity of the substance \((c)\).

Thus,

\[
Q = mc\Delta T
\]

Sample problem 1.3

(a) How much energy is needed to heat 8.0 L (about 8.0 kg) of water from a room temperature of 15 °C to 85 °C (just right for washing dishes)?

(b) A chef pours 200 g of cold water with a temperature of 15 °C into a hot aluminium saucepan with a mass of 250 g and a temperature of 120 °C. What will be the common temperature of the water and saucepan when thermal equilibrium is reached? Assume that no energy is transferred to or from the surroundings.

**Solution:**

(a) \( Q = mc\Delta T \)

where

\[
c = 4200 \text{ J kg}^{-1} \text{ K}^{-1} \quad \text{(from table 1.2)}
\]

\[
m = 8.0 \text{ kg}
\]

\[
\Delta T = 70 \text{ K} \quad \text{(same change as 70 °C)}
\]

Therefore,

\[
Q = 8.0 \text{ kg} \times 4200 \text{ J kg}^{-1} \text{ K}^{-1} \times 70 \text{ K} \quad \text{(substituting data)}
\]

\[
= 2352000 \text{ J} \quad \text{(solving)}
\]

\[
= 2.35 \times 10^6 \text{ J} \quad \text{(using the most appropriate units)}
\]

The energy needed is best expressed as 2400 kJ.

(b) The solution to this question relies on the following three factors.

1. Energy is transferred from the saucepan into the water until both the saucepan and the water reach the same temperature \((T_f °C)\).
2. The amount of internal energy \((Q_w)\) gained by the water will be the same as the amount of internal energy lost by the saucepan \((Q_s)\).
3. The internal energy gained or lost can be expressed as \(mc\Delta T\). \((\Delta T\) can be expressed in K or °C since change in temperature is the same in both units.)

Therefore,

\[
Q_w = Q_s
\]

\[
m_w c_w \Delta T_w = m_s c_s \Delta T_s
\]

where

change in temperature of the water, \(\Delta T_w = T_f - 15 °C\)

change in temperature of the saucepan, \(\Delta T_s = 120 °C - T_i\)

\[
0.200 \text{ g} \times 4200 \text{ J kg}^{-1} \text{ C}^{-1} \times (T_f - 15 °C) = 0.250 \text{ g} \times 900 \text{ J kg}^{-1} \text{ C}^{-1} \times (120 - T_i)
\]

(substituting data)

\[
840T_f - 12 \times 600 = 27 \times 000 - 225T_i
\]

(simplifying and expanding brackets)

\[
1065T_f = 39 \times 600
\]

\[
T_f = 37 °C
\]

(solving)

The saucepan and water will reach a common temperature of 37 °C.
This example is a good illustration of the implications of a high specific heat capacity. Even though there was a smaller mass of water than aluminium, the final temperature was closer to the original water temperature than the original aluminium temperature.

**Revision question 1.4**

How much energy is need to increase the temperature of your body by 1 °C?

**AS A MATTER OF FACT**

Eating a hot pie can be a health hazard! The temperature of the pastry and filling of a hot pie are the same. Thermal equilibrium has been reached. So why can you bite into a pie that seems cool enough to eat and be burnt by the filling?

The reason is that the filling is mostly water, while the pastry is mostly air. When your mouth surrounds that tasty pie, energy is transferred from the pie to your mouth. Each gram of water in the filling releases about 4 J of energy into your mouth for every 1 °C lost (since the specific heat of water is 4200 J kg\(^{-1}\) K\(^{-1}\)). Each gram of air in the pastry releases only about 1 J of energy into your mouth for every 1 °C lost (since the specific heat of air is 1000 J kg\(^{-1}\) K\(^{-1}\)). Gram for gram, the filling transfers four times more energy into your mouth than the pastry.

**Latent heat and the kinetic particle model of matter**

In order for a substance to melt or evaporate, energy must be added. During the process of melting or evaporating, the temperature of the substance does not increase. The energy added while the state is changing is called latent heat. The word latent is used because it means ‘hidden.’ The usual evidence of heating, a change in temperature, is not observed.

Similarly, when substances freeze or condense, energy must be released. However, during the process of changing state, there is no decrease in temperature accompanying the loss of internal energy.

In simple terms, the energy transferred to or from a substance during melting, evaporating, freezing or condensing is used to change the state rather than to change the temperature.

During a change of state, internal energy is gained or lost from the substance. Recall, however, that the internal energy includes the random kinetic and potential energy of the particles in the substance. The random translational kinetic energy of particles determines the temperature.

When a substance being heated reaches its melting point, the incoming energy increases the potential energy of the particles rather than the random translational kinetic energy of the particles. After the substance has melted completely, the incoming energy increases the kinetic energy of the particles again. When the substance is being cooled, the internal energy lost on reaching the melting (or freezing) point is potential energy. The temperature does not decrease until the substance has completely solidified.

The same process occurs at the boiling point of a substance. While evaporation or condensation takes place, the temperature of the substance does not change. The energy being gained or lost is latent heat, ‘hidden’ as changes in internal potential energy take place.
The specific latent heat of fusion is the quantity of energy required to change 1 kilogram of a substance from a solid to a liquid without a change in temperature.

The specific latent heat of vaporisation is the quantity of energy required to change 1 kilogram of a substance from a liquid to a gas without a change in temperature.

**Specific latent heat of fusion**

The specific latent heat of fusion is the quantity of energy required to change 1 kilogram of a substance from a solid to a liquid without a change in temperature. Note that the same quantity of energy is lost without a change in temperature during the change from a liquid to a solid. The specific latent heat of fusion of water is 334 kJ kg\(^{-1}\).

**Specific latent heat of vaporisation**

The specific latent heat of vaporisation is the quantity of energy required to change 1 kilogram of a substance from a liquid to a gas without a change in temperature. Note that the same quantity of energy is lost without a change in temperature during the change from a gas to a liquid. The specific latent heat of vaporisation of water is \(2.3 \times 10^3\) kJ kg\(^{-1}\).

<table>
<thead>
<tr>
<th>Substance</th>
<th>Specific latent heat of fusion (J kg(^{-1}))</th>
<th>Specific latent heat of vaporisation (J kg(^{-1}))</th>
</tr>
</thead>
<tbody>
<tr>
<td>Water</td>
<td>(3.3 \times 10^5)</td>
<td>(2.3 \times 10^6)</td>
</tr>
<tr>
<td>Oxygen</td>
<td>(6.9 \times 10^3)</td>
<td>(1.1 \times 10^5)</td>
</tr>
<tr>
<td>Sodium chloride</td>
<td>(4.9 \times 10^5)</td>
<td>(2.9 \times 10^6)</td>
</tr>
<tr>
<td>Aluminium</td>
<td>(2.2 \times 10^3)</td>
<td>(1.7 \times 10^4)</td>
</tr>
<tr>
<td>Iron</td>
<td>(2.8 \times 10^5)</td>
<td>(6.3 \times 10^6)</td>
</tr>
</tbody>
</table>

The graph on the left shows how the temperature of water increases as it is heated at a constant rate. During the interval BC, the temperature is not increasing. The water is changing state. The energy transferred to the water is not increasing the random translational kinetic energy of water particles. Note that the gradient of the section AB is considerably less than the gradient of the section CD. What difference in the properties of water and steam does this reflect?

Algebraically, the quantity of energy, \(Q\), required to change the state of a substance without a change in temperature can be expressed as:

\[ Q = mL \]

where

- \(m\) = mass of the substance
- \(L\) = specific latent heat of fusion or vaporisation.
Evaporation
Your skin is not completely watertight, which allows water from the skin and tissues beneath it to evaporate. The latent heat of vaporisation required for the water to change state from liquid to gas is obtained from the body, reducing its temperature. Evaporation of water from the mouth and lungs also takes place during the process of breathing. Even without sweating, the energy used to evaporate water in the body accounts for about 17% of the total heat transfer from the body to the environment.

Water evaporates even though its temperature is well below its boiling point. The temperature is dependent on the average translational kinetic energy of the water molecules. Those molecules with a kinetic energy greater than average will be moving faster than the others. Some of them will be moving fast enough to break the bonds holding them to the water and escape from the liquid surface. The escaping molecules obtained their additional energy from the rest of the liquid water, thus reducing its temperature.

**As a Matter of Fact**
A burn caused by steam at 100 °C is considerably more serious than a burn caused by the same mass of boiling water. Each kilogram of hot steam transfers 2600 kJ of energy to your skin as it condenses to water at 100 °C. Each kilogram of newly condensed steam then transfers 4.2 kJ for each °C drop in temperature as it cools to your body temperature of about 37 °C. That’s about 265 kJ. The total quantity of energy transferred by each kilogram of steam is therefore about 2865 kJ. A kilogram of boiling water would transfer 265 kJ of energy as it cooled to your body temperature.
Chapter review

Summary

- A thermometer measures temperature, and various properties of materials can be used to make one.
- There are different temperature scales, with Celsius and Kelvin being the common ones. Temperatures in one scale can be converted to any other.
- The kinetic particle model of matter explains heat phenomena.
- Internal energy is the energy associated with the random movement of molecules and it comes in many forms, including translational kinetic energy, rotational and vibrational kinetic energy, and potential energy.
- Temperature is a measure of the average translational kinetic energy of the atoms and molecules in a substance.
- Objects at different temperatures, if placed in contact, will reach a common temperature. This process is called thermal equilibrium and is described as the Zeroth Law of Thermodynamics.
- The First Law of Thermodynamics states that if energy is transferred to or from a system, then the total energy must be conserved, with any changes in the internal energy of the system given by \( \Delta U = Q - W \), where \( U \) is the change in internal energy, \( Q \) is the heat added to the system and \( W \) is the work done by the system.
- The specific heat capacity, \( c \), of a substance is the amount of energy required to increase the temperature of 1 kg of the substance by 1 °C.
- When substances of different specific heat capacities and different temperatures are mixed, the final temperature can be determined by using the conservation of energy and the relationship \( Q = mc\Delta T \) for each substance.
- The latent heat, \( L \), of a substance is the amount of energy required to change the state from solid to liquid or liquid to gas or vice versa of the substance. For a substance of mass \( m \) kg, the energy required is given by \( Q = mL \).
- Evaporation of a liquid occurs because some of the fastest-moving particles have sufficient energy and speed to break free of the surface. The removal of these particles increases the overall average kinetic energy of the remaining particles and consequently the temperature of the liquid.

Questions

Temperature conversion

1. Why is the Celsius scale of temperature commonly used rather than the Kelvin scale?
2. What is the main advantage of an absolute scale of temperature?

3. Estimate each of the following temperatures in Kelvin:
   (a) the maximum temperature in Melbourne on a hot summer’s day
   (b) the minimum temperature in Melbourne on a cold, frosty winter’s morning
   (c) the current room temperature
   (d) the temperature of cold tap water
   (e) the boiling point of water

4. The temperature of very cold water in a small test tube is measured with a large mercury-in-glass thermometer. The temperature measured is unexpectedly high. Suggest a reason why this might be the case.

Particle model of matter

5. James Joule showed that mechanical energy could be transformed into the internal energy of a substance or object. The temperature of a nail, for example, can be raised by hitting it with a hammer. List as many examples as you can of the use of mechanical energy to increase the temperature of a substance or object.

6. Explain in terms of the kinetic particle model why a red-hot pin dropped into a cup of water has less effect on the water’s temperature than a red-hot nail dropped into the same cup of water.

7. If today’s maximum temperature was 14 °C and tomorrow’s maximum temperature is expected to be 28 °C, will tomorrow be twice as hot? Explain your answer.

8. Explain why energy is transferred from your body into the cold sea while swimming even though you have less internal energy than the surrounding cold water.

9. Why can’t you put your hand on your own forehead to estimate your body temperature?

10. It is said that thermometers indirectly measure the temperature of an object by measuring their own temperature. Explain this statement by referring to the concept of thermal equilibrium.

11. Adam says that ‘A thermometer measures the average temperature between itself and the object it is measuring,’ while Bob says that ‘A thermometer directly measures the temperature of the object.’ Explain why each is wrong.

First Law of Thermodynamics

12. For each of the following, calculate the values of \( Q \), \( W \) and \( \Delta U \) and indicate whether the temperature increases, decreases or stays the same.
   (a) A gas in a fixed container is heated by 500 J.
(b) A gas in a container with a flexible lid is cooled by ice with 250 J of energy extracted.
(c) A gas in a container with a plunger is squashed by a heavy mass moving down, losing 150 J of gravitational potential energy.
(d) A stretched rubber band at room temperature with 5 J of stored energy is released.

13. A can filled with a high-pressure gas has a balloon attached over the top. What happens to the temperature of the gas inside the can as you allow the gas to expand into the balloon?

14. In the figure above, two beakers are filled with the same gas. A plunger is fitted so that no gas escapes; friction is negligible between the plunger and the beaker walls.

(a) The block is removed from each plunger, the plunger moves upward.
(b) In each case, does the gas do work or is work done on the gas? Explain your reasoning in a few sentences.
(c) Is there a larger transfer of thermal energy as heat between Gas A and the surroundings or between Gas B and the surroundings? Explain your reasoning in a few sentences. Draw an arrow on each figure indicating the direction of thermal energy flow.
(d) For the expansion of Gas A, how do the work and heat involved in this process affect the internal energy of the gas? Explain your reasoning in a few sentences.
(e) For the expansion of Gas B, how do the work and heat involved in this process affect the internal energy of the gas? Explain your reasoning in a few sentences.

15. A barbecue uses gas from a gas bottle as its energy source. After the BBQ has been running a while, ice is noticed around the top of the gas bottle. Explain the physics principles behind this observation.

16. Two insulated containers are connected by a valve. The valve is closed. One container is filled with gas, the other is a vacuum. The valve is opened. Is there any change in temperature? Is there any change in the internal energy? What are the values of $Q$ and $W$ in this situation?

17. Consider these three scenarios then complete the table below, using either 0, − or +.
   A. An insulated container, such as a thermos flask, has a piston that can be moved up and down without letting the air out. The piston is pushed down.
   B. A metal tin with a lid is heated.
   C. A metal tin with a sliding lid and a mass on top is heated.

<table>
<thead>
<tr>
<th></th>
<th>A</th>
<th>B</th>
<th>C</th>
</tr>
</thead>
<tbody>
<tr>
<td>Heat (+ is in)</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Work (+ is out)</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>$\Delta U$</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

18. The same hotplate is used to heat 50 g of ethylene glycol (used in car antifreeze) and 50 g of cooking oil. Both substances are heated for 2 minutes. Use the data in the table above to determine:
   (a) which liquid needs more energy to raise its temperature by 1 °C
   (b) which liquid will experience the greater increase in temperature.

19. The quantity of energy needed to increase the temperature of a substance is directly proportional to the mass, specific heat capacity and the change in temperature of the substance. If 200 kJ is used to increase the temperature of a particular quantity of a substance, how much energy would be needed to bring:
   (a) twice as much of the substance through the same change in temperature?
(b) three times as much of the substance through a temperature change twice as great?

20. Use the table on page 19 to answer the following questions.
   (a) Why is the specific heat capacity of the human body so high?
   (b) Why is the specific heat capacity of desert sand so much lower than that of fertile topsoil?
   (c) When heating water to boiling point in a saucepan, some of the energy transferred from the hotplate is used to increase the temperature of the saucepan. Which would you expect to gain the most energy from the hotplate: an aluminium, copper or steel saucepan?
   (d) Make some general comments about the order of substances listed in the table on page 19.

21. An 800 g rubber hot-water bottle that has been stored at a room temperature of 15 °C is filled with 1.5 kg of water at a temperature of 80 °C. Before being placed in a cold bed, thermal equilibrium between the rubber and water is reached. What is the common temperature of the rubber and water at this time? (Assume that no energy is lost to the surroundings. The specific heat capacity of rubber is 1700 J kg⁻¹ K⁻¹. The specific heat capacity of water is 4200 J kg⁻¹ K⁻¹.)

Latent heat

22. Use the data below to determine the quantity of energy needed to evaporate 500 g of water without a change in temperature.

<table>
<thead>
<tr>
<th>Substance</th>
<th>Specific latent heat of fusion (J Kg⁻¹)</th>
<th>Specific latent heat of vaporisation (J Kg⁻¹)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Water</td>
<td>3.3 × 10⁵</td>
<td>2.3 × 10⁶</td>
</tr>
</tbody>
</table>

23. The graph below shows the heating curve obtained when 500 g of candle wax in solid form was heated from room temperature in a beaker of boiling water.

(a) What is the boiling point of the candle wax?
(b) During the interval BC, there is no increase in temperature even though heating continued. What was the energy transferred to the candle wax being used for during this interval?
(c) In which state of matter was the candle wax during the interval CD?
(d) Use the heating curve to determine the latent heat of fusion of candle wax.
(e) Which is higher: the specific heat capacity of solid candle wax or the specific heat capacity of liquid candle wax? Explain your answer.
(f) Explain in terms of the kinetic particle model what is happening during the interval DE.

24. How much energy does it take to completely convert 2 kg of ice at −5 °C into steam at 100 °C? Assume no energy loss to the surroundings.

25. How much ice at 0 °C could be melted with 1 kg of steam at 100 °C, assuming no loss of energy to the surroundings? Use the specific latent heat values quoted in table 1.4 on page 16. The specific heat capacity of water is 4200 J kg⁻¹ K⁻¹.

26. Explain why vegetables cook faster by being steamed than boiled.

27. Why are burns caused by steam more serious than those caused by boiling water?

28. In hot weather, sweat evaporates from the skin. Where does the energy required to evaporate the sweat come from?

29. Explain the importance of keeping a lid on a simmering saucepan of water in terms of latent heat of vaporisation.

30. Explain in terms of the kinetic particle model why you can put your hand safely in a 300 °C oven for a few seconds, while if you touch a metal tray in the same oven your hand will be burned.

31. How does the evaporation of water cause a reduction in the temperature of the surrounding air?

32. Give two reasons why you feel cooler when the wind is blowing than you would in still air at the same temperature.

33. In humid weather, evaporation of perspiration takes place as it does in dry weather. However, the cooling effect is greatly reduced. Why?