UNIT 04

Thermal Physics

CHAPTER 10

Heat and Temperature

10.1 HEAT AND HUMANS

Heat, and the lack of heat, have been important to humankind since the earliest times. Heat from the Sun and the cold of winter have always affected living conditions. From early childhood, hotness and coldness are two of the first feelings children encounter. Children experience these conditions early in life by sucking an iceblock or walking on a hot road. In fact it is the temperature change that initiates breathing when a child is born. The industrial revolution was based on heat and the generation of mechanical power from heat. Many industries in modern-day life are solely centred around the production of heat or the purposeful removal of heat, for example refrigeration, airconditioners and pot-belly stoves.

NEI Activity 10.1 HOT SPOTS

People have always wondered about the extremes of temperature. Consult the *Guinness Book of Records* or the Internet to find:

- the highest and lowest recorded human body temperature
- the places on earth that recorded the highest and lowest temperature
- the highest and lowest temperature ever achieved on Earth.

But what is heat?

10.2 HEAT AND TEMPERATURE

Up to the eighteenth century heat was regarded as some sort of invisible fluid, a 'caloric fluid' that bodies possessed. Hot bodies, it was believed, contained more of this fluid than cold bodies. When a body was warmed this caloric fluid was added to the body. But this did not explain why two ice cubes melted when they were rubbed together.

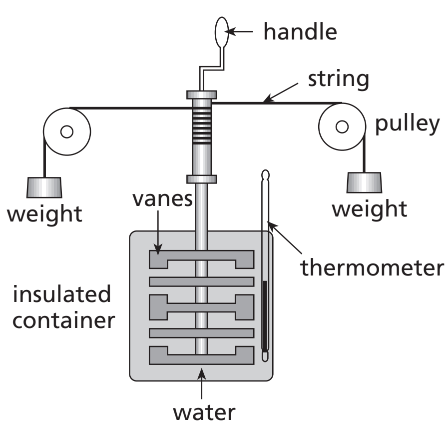


Figure 10.1
A schematic diagram of Joule's apparatus used to investigate the relationship between mechanical and thermal energy.

English scientist **James Joule** (1818–89) was one of the first scientists to show that heat was a form of energy. He performed an experiment in which falling lead weights turned paddles in water. The work done by these weights caused the water to heat up. He showed that mechanical energy can be converted into heat. Joule concluded that heat is a form of energy, but what form does it take?

Consider a student throwing a ball. (See Figure 10.2.)

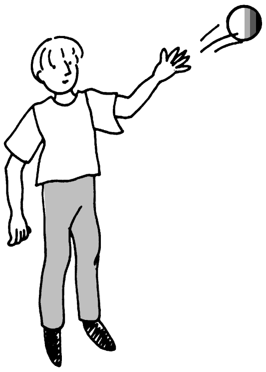


Figure 10.2
The energy contained in a ball in flight due to its motion and height.

What energy does the ball possess?

Some might say it has kinetic energy because of its motion. Others might say it has potential energy due to its height above the ground. Some would say both. Is that all the energy it possesses?

Now consider the same student throwing a box full of bees (Figure 10.3). What is the total energy of the container?

Figure 10.3
The energy contained in a box of bees, including motion of the bees.

It is easily seen that the container possesses both kinetic energy due to the motion of the box, and potential energy due to its height above the ground. However, it also possesses the kinetic energy the bees themselves might have.

Up to this time we have considered the bulk energy of objects and not the internal energy the particles in the objects might possess. All objects contain particles, atoms and/or molecules, and it is the motion of these particles that makes up the internal energy of the object. This motion can be vibrational kinetic energy, as in solids; or rotational and translational kinetic energy, as in fluids. The particles of matter also possess many forms of potential energy in the bonds that hold particles together, as well as that stored in the nucleus of the atoms. (See Figures 10.4 **(a)**, **(b)** and **(c)**.)

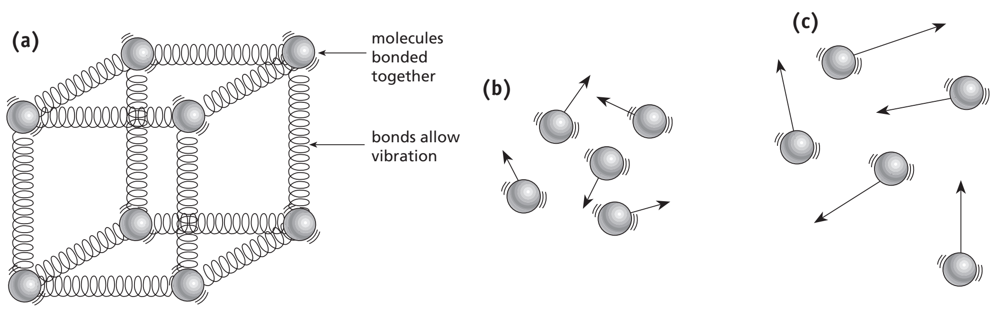


Figure 10.4
(a) A model of the movement of the molecules in a solid. The springs represent the molecular bonds. **(b)** A representation of the molecules of a liquid. They vibrate and are able to flow over one another. **(c)** The molecules of a gas have greater freedom. They have rapid straight line motion.

The sum of the kinetic and potential energies of all the particles is called the **internal** or **thermal energy** of the object. Heating is the term used when some of the thermal energy is transferred from hot objects to cold objects as in the case of a hot spoon being placed in cool water. The term heat is used to describe the internal energy transferred through this heating process. The study of these energy transfers is called **thermodynamics** (from the Greek *thermos* meaning 'heat', and *dynamis* meaning 'powerful').

It would be impossible to measure the motion of all the particles within a substance because of the number of particles and the great variation in speeds of the particles. Figure 10.5 indicates the variation of molecular speeds of a gas at various temperatures.

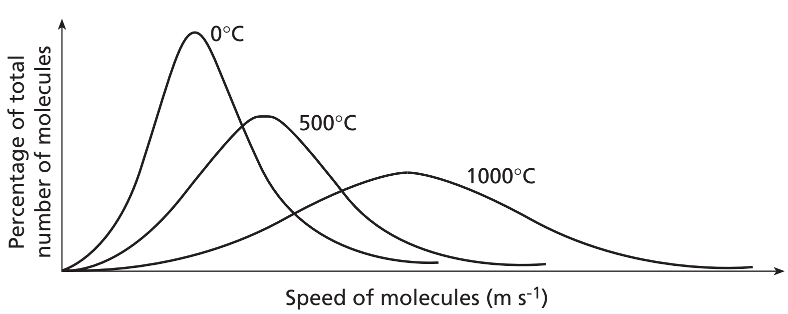


Figure 10.5
The distribution of speeds of the molecules of a gas at various temperatures.

However, as objects gain heat and become hotter, the particles move faster. Temperature is a measure of the average kinetic energy of the particles of the object. Changing the potential energy of a substance without changing the average kinetic energy of its molecules does not change the temperature of the substance. This occurs when a change of state occurs. That is, when a substance changes from a solid to a liquid, or a liquid to a gas, or vice versa. A common misconception is that heat and temperature are the same, which is not the case.

10.3 THERMAL EQUILIBRIUM

Everyone has observed that when a cool spoon is placed in a hot cup of coffee it eventually becomes hot — as hot as the coffee. Or when you use a thermometer to measure a person's temperature the thermometer becomes as warm as the person whose temperature is being measured. How does this occur?

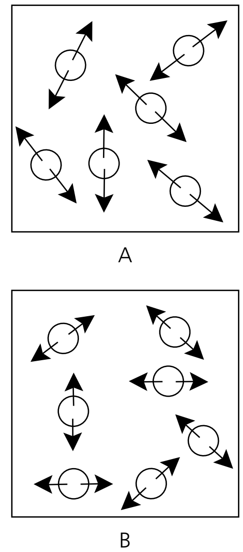
Consider a closed system (a system where there are no energy losses to the environment) in which a hot object A is in contact with a cooler one B. (See Figure 10.6.)

Because A is hot it contains more thermal energy than B, and its molecules have more potential and kinetic energy. The molecules move faster in object A than they do in object B. When A and B are placed in contact the molecules of A collide with the molecules of B, transferring kinetic energy to them. This causes the molecules of B to vibrate further apart, thus increasing object B's potential energy. Object B's thermal energy has increased. At the same time the molecules of A have slowed down and vibrate closer together, thus decreasing A's kinetic and potential energies. Object A has lost thermal energy. Thermal equilibrium is reached when the energy given to B equals the energy B is giving back to A. As the law of conservation of energy is true for all forms of energy:

heat lost by object A = heat gained by object B

So when you use a thermometer to measure a child's temperature, molecules of the child in contact with the thermometer jostle the molecules of the glass, which in turn jostle the molecules of the mercury in the thermometer. The mercury expands, indicating the temperature on an appropriate scale of temperature.

Figure 10.6 The flow of thermal energy from the molecules of a hot to those of a cold body.



Activity 10.2 FREEZING

EI

NEI

1 Find out which freezes first — a cup of hot water or a cup of cool water placed in the freezer.

2 Do pets have the same body temperature as humans?

Questions

1 Which has more thermal energy: a cup of water at 100°C or a bath full of water at 40°C? Why?

2 Steam at 100°C will give you much more severe burns than water at 100°C.
(a) In which one are the molecules moving the faster?
(b) In which one do the molecules have greater potential energy?
(c) Why are steam burns more severe?

3 **(a)** What is the name given to the internal energy of a substance?
(b) What form(s) of energy does this involve?

4 If James Joule did 100 J of work on several quantities of water — 100 mL, 300 mL and 500 mL:
(a) which sample would gain the most thermal energy? Why?
(b) which sample's temperature would increase the most? Why?

NOVEL CHALLENGE

During the Second World War, Nazi scientists threw many prisoners overboard into the freezing waters of the North Sea to see how fast their body temperature dropped and how long it would take for them to die. Today, such data are needed by ocean rescue researchers to help to develop safety devices for ocean users.
Should we use this despicable 'Nazi science'? Develop an argument for or against its use.

NOVEL CHALLENGE

The table below shows the effects of changes to body temperature:

T (°C)	*Effect*
37.0 ± 1	normal oral
35	shivering
34	slurred speech
33	hallucinations
32	shivering stops
30	unconsciousness
26	appears dead

Death is defined as a failure to revive on rewarming above 32°C. When people freeze to death in cold water it has been reported that they do not seem to be in pain as they die. They often seem relaxed. *What could be happening here?*

10.4 MEASURING TEMPERATURE

Measuring temperature requires the use of some property of a substance that changes proportionally with increase in temperature. The property that most temperature measuring instruments use is expansion and contraction. This is the property used in most thermometers. The most common is mercury-in-glass or alcohol-in-glass thermometers that have

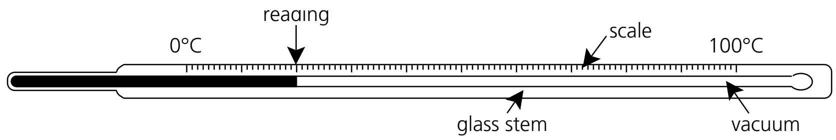


Figure 10.7
The liquid-in-glass thermometer.

been calibrated to indicate the temperature. As temperature increases, the mercury or alcohol expands up a fine tube in the glass thermometer. The markings on the thermometer depend on the scale used. (See Figure 10.7.)

Throughout history scientists made up their own scales to measure temperature. Sir Isaac Newton made up a temperature scale where the freezing point of water was 0 and normal body temperature was 12.

NOVEL CHALLENGE

At what temperature will °C and °F readings be the same? Could the Kelvin temperature reading ever be the same as °C or °F reading?

The Fahrenheit scale

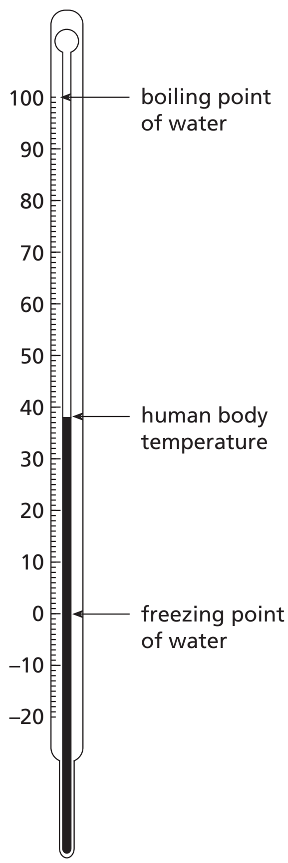
A German physicist, **Gabriel Fahrenheit** (1686–1736), developed a liquid-in-glass thermometer and a temperature scale that took the freezing point of an ice and salt mixture to be 0°F. He took the freezing point of pure water as 32°F and normal body temperature to be 96°F. The boiling point of water is then 212°F. This scale is no longer used in Australia but is still in use in several other countries such as the USA, UK and Canada. The conversion is $°C = (°F - 32) \times \frac{5}{9}$.

The Celsius scale

An easier decimal scale was invented by a Swedish astronomer, **Anders Celsius** (1701–44). On the Celsius scale, also called the centigrade scale, the freezing point of pure water is 0°C and the boiling point is 100°C. Interestingly, he originally took the freezing point to be 100°C and boiling point to be 0°C, but this was changed in the first year. This is the main scale used in measuring body temperature. (See Figure 10.8.)

Did you know that the body temperature of a baby is higher than that of an adult? Did you know also that aspirin is used to lower the body temperature?

Figure 10.8
A Celsius thermometer, a common thermometer used in the laboratory and in the home.



The Kelvin/absolute temperature scale

The two previous scales are relative scales. That is, zero degree on either scale does not mean that this is the lowest temperature obtainable. Since temperature is a measure of the average kinetic energy of the particles, 0°C does not mean that all particle motion has stopped. Then at what temperature does all motion stop? This point would be the true limit of coldness and would produce an absolute zero temperature. **Lord Kelvin** (1824–1907) suggested this temperature was –273.15°C.

When a sample of gas of constant volume is heated its pressure varies with Celsius temperature, as shown in Figure 10.9. Extrapolation of this graph suggests that, at –273.15°C, the pressure becomes zero and therefore all particle motion stops. This is because pressure is caused by particles colliding with the container walls and if there is no motion there are no collisions and therefore no pressure. This point is called absolute zero on the Kelvin scale of temperature. However, one degree on the Kelvin scale is equal in magnitude to one degree on the Celsius scale.

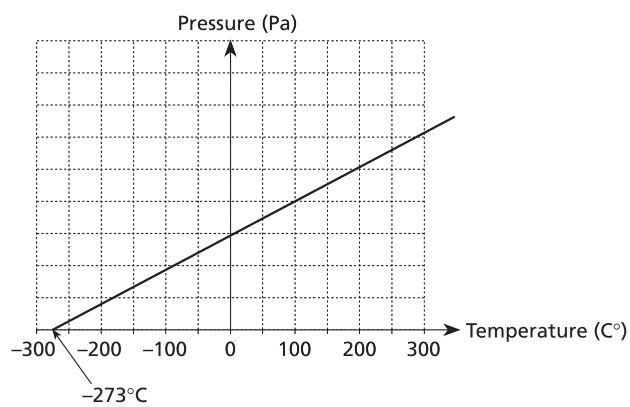


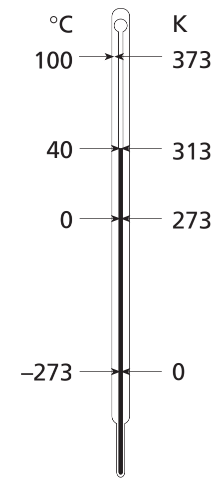
Figure 10.9
The relationship between temperature and pressure, and the establishment of absolute zero.

Therefore changing Celsius temperature to Kelvin temperature simply requires the addition of 273 (to three significant figures) to the Celsius value. Figure 10.10 shows a comparison between the two scales.

Kelvin temperature = Celsius temperature + 273

K = °C + 273

Figure 10.10
A comparison of temperature in degrees Celsius and Kelvin.



Example 1

Convert 50°C to kelvins.

Solution

$$\begin{aligned} K &= °C + 273 \\ &= 50°C + 273 \\ &= 323\ K \end{aligned}$$

Example 2

Convert 486 K to °C.

Solution

$$\begin{aligned} K &= °C + 273 \\ 486 &= °C + 273 \\ 486 - 273 &= °C \\ &= 213°C \end{aligned}$$

Questions

5 Convert the following temperatures to K: **(a)** 20°C; **(b)** –150°C; **(c)** 520°C; **(d)** –72°C; **(e)** –300°C.

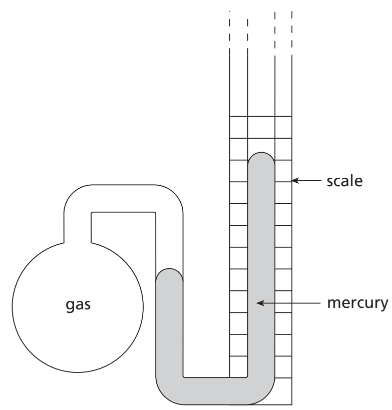
6 Convert the following temperatures to °C: **(a)** 50 K; **(b)** 278 K; **(c)** 1000 K; **(d)** –50 K.

Other types of thermometers

Even though the liquid-in-glass thermometers are the most widely used in science and in general, they have their limitations. This is mainly due to the liquid freezing or boiling. Alcohol-in-glass thermometers can be used between –100°C and 80°C. Mercury-in-glass thermometers have an operating range of –40°C to 360°C. Glass is also fragile, and mercury is toxic to the body and the environment.

Gas thermometers rely on the expansion of gas. Since change in temperature is proportional to the change in volume of a gas, the expansion of a gas can be calibrated to measure temperature.

Figure 10.11
A constant volume gas thermometer — a thermometer that relies on the relationship between temperature and pressure.



Resistance thermometers use the fact that electric current in wires decreases as temperature rises.

Thermocouples consist of two wires made of different metals. The wires are made into a loop including a voltmeter. One end is kept at a reference temperature and the other end is used as a probe. (See Figure 10.12.) When this probe is placed in a substance to be measured the voltage produced is proportional to the difference in temperature between the two ends. Thermocouples can be used to measure temperature over a wide range.

Bimetallic strips rely on the different expansion rates of two different metals. When heated one metal expands more than the other, causing bending and movement of a pointer across a scale. These have a wide working range.

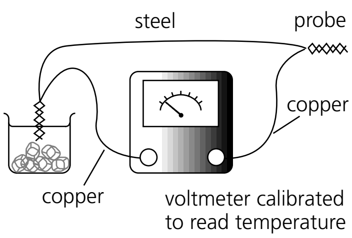
In *liquid crystal thermometers* numbers on a scale are made of different crystalline chemicals. As temperature increases these chemicals change their crystalline structure, which results in colour changes. These are not very accurate.

Pyrometers measure the radiation given off by objects. The characteristic of the radiation changes with temperature. Infrared pyrometers can measure temperature from –20°C to 1500°C. Body temperature is routinely monitored in clinical settings with infrared ear thermometers, which measure the infrared energy emitted from the patient's eardrum in a calibrated length of time. A short tube with a protective sleeve is inserted into the ear, and a shutter is opened to allow radiation from the tympanic membrane to fall on an infrared detector for 0.1 to 0.3 seconds. The device beeps when data collection is completed and a readout of temperature is produced on a liquid crystal display.

Thermistors are semiconductor devices that change their resistance with change in temperature. When these devices are heated their resistance decreases and more current flows. The current is measured on an ammeter, which is calibrated to read temperature.

SPECIFIC HEAT CAPACITY 10.5

Figure 10.12
A thermocouple — a thermometer consisting of two dissimilar metals. The difference in the temperatures of the two ends produces a voltage. This can be calibrated to 'read' temperature.



Do all objects increase their temperature at the same rate as they are heated?

Because of the great variation in molecular structure and bonds that exist between atoms in different substances, energy put into different substances does not result in the same temperature rises. For example, when walking along a beach on a hot day the sand is a lot hotter than the grass or puddles of water. This is because sand only requires 880 J of heat energy from the sun to raise 1 kg of sand by 1°C. Water requires 4200 J.

This property is called the specific heat capacity, *c*, of the substance. It is defined as the amount of energy required to raise the temperature of 1 kg of a substance by 1°C or by 1 K (a change of 1°C is equivalent to a change of 1 K).

Heat capacity is the term used to describe the amount of heat an object contains. Different materials of the same mass and at the same temperature contain different amounts of heat energy. That is, they require different amounts of heat to heat them up and they give out different amounts of energy in cooling down. For example, which contains the more heat — a BBQ plate at 120°C or a coin at the same temperature?

The specific heat capacity of some common substances is given in Table 10.1.

Figure 10.13
A bimetallic strip thermometer. The difference in expansion rates between the two different metals causes a pointer to move across a scale.

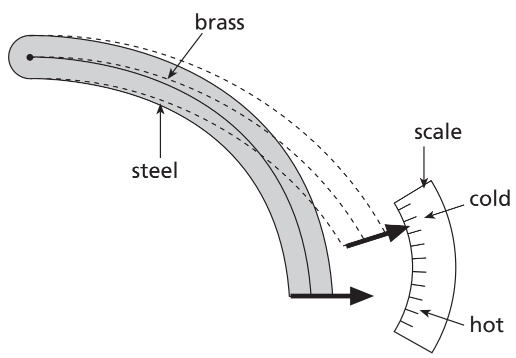


Table 10.1 SPECIFIC HEATS OF SOME COMMON SUBSTANCES

SUBSTANCE	SPECIFIC HEAT CAPACITY c (J kg^{-1} K^{-1})
Lead	130
Mercury	140
Copper	390
Iron and steel	460
Glass	664
Sodium chloride	880
Sand	880
Aluminium	900
Wood	1700
Steam	2020
Ice	2100
Paraffin	2200
Honey	2370
Alcohol	2450
Methylated spirits	2500
Water	4200

Which substance has the highest specific heat capacity?

This is important. Because water has a high specific heat capacity compared with other substances, interesting uses are made of water. Water is used in cooling systems in motor cars. One kilogram of water can take away 4200 J of heat from the engine of a car before its temperature changes by 1°C. What would happen if alcohol was used? The ocean's temperature changes by small amounts compared with the land mass from winter to summer and from day to night. Humans are composed of approximately 70% water, which results in humans responding less to external temperature changes than if they were composed of different materials. Consider how Superman (the man of steel) would be affected by changes in the external temperature.

A useful equation to determine the specific heat of a substance is:

$$Q = mc\Delta T$$

where c is the specific heat capacity in J kg^{-1} °C^{-1}, Q is the quantity of heat in J, m is the mass of the object in kg, ΔT is the change in temperature in °C.
Note: if ΔT is negative then this is the quantity of heat given off by an object.

Example 1

If it takes 4000 J of heat energy to raise the temperature of a 2 kg object by 10°C, what is the specific heat capacity of the object?

Solution

$$\begin{aligned} Q &= mc\Delta T \\ 4000 &= 2 \times c \times 10 \\ c &= 2.00 \times 10^2 \text{ J kg}^{-1}\text{ K}^{-1} \end{aligned}$$

Example 2

How much heat is required to bring a saucepan containing 500 mL of water at 20°C to boiling point? (1 L of water has a mass of 1 kg.)

Solution

$$\begin{aligned} Q &= mc\Delta T \\ &= 0.5 \times 4200 \times (100 - 20) \\ &= 0.5 \times 4200 \times 80 \\ &= 1.68 \times 10^5 \text{ J} \end{aligned}$$

Example 3

If it takes 2850 J of heat to raise the temperature of a 1.5 kg object by 5°C, what substance could the object be made of?

Solution

$$\begin{aligned} Q &= mc\Delta T \\ 2850 &= 1.5 \times c \times 5 \\ c &= 3.80 \times 10^2 \text{ J kg}^{-1}\text{ K}^{-1} \end{aligned}$$

The substance is made of copper (from Table 10.1).

Example 4

A 2 kg block of iron at 25°C is heated by placing it in hot water. If it obtains 4500 J of heat from the water, what is the final temperature of the iron?

Solution

$$\begin{aligned} Q &= mc\Delta T \\ 4500 &= 2 \times 450 \times (T_f - T_i) \\ 4500 &= 900 \times (T_f - 25) \\ T_f &= 50 + 25 \\ T_f &= 75\text{°C} \end{aligned}$$

NOVEL CHALLENGE

When you eat ice cubes, your body uses up energy to melt them and warm them up to body temperature.
Does this mean eating ice would help you lose weight? Could you say they have negative joules?

NOVEL CHALLENGE

In 1700, Dr Charles Blagden took some friends, a dog and a raw beefsteak into a room at 127°C for $\frac{3}{4}$ hour. They all came out unharmed — except for the steak, which was cooked. Why?

NOVEL CHALLENGE

Imagine that you added equal volumes of water and oil to separate beakers and placed them on a hotplate. And after 5 minutes the water began boiling.
Which liquid would be at the higher temperature? Which liquid would have the greater total thermal energy?

Questions

7 Calculate the heat energy absorbed when 1.5 kg of paraffin is raised from 15°C to 50°C.

8 How much heat could be absorbed by a 3 kg block of ice at −10°C before it reaches its melting point?

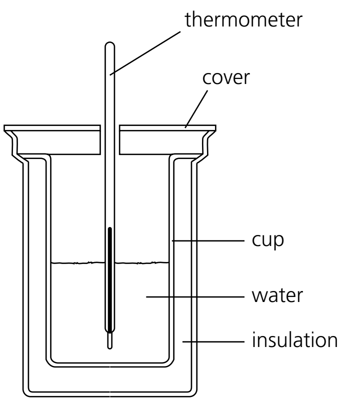
9 A liquid is heated in a beaker. If it is found that it takes 7500 J of heat to increase the temperature of 500 g of the liquid by 6°C, what could the liquid be?

CALORIMETRY 10.6

If two substances are placed together in a closed system, that is, one where no energy can escape to the surroundings, then the heat energy lost by one object in the exchange is equal to the heat energy gained by the other. For example, when a cool teaspoon is placed in a hot cup of coffee the heat lost by the coffee and the cup is equal to the heat gained by the spoon.

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

Figure 10.14
A calorimeter — a device used to measure heat exchanges effectively by minimising heat losses to the environment.



This principle is just an extension of the law of conservation of energy — energy is not lost or gained, just transferred or transformed.

In practice there is always some heat lost unless insulation is ideal. However, heat losses to the surroundings can be minimised if experiments are carried out quickly. Scientific experiments also use calorimeters (Figure 10.14), which have good insulation to limit the loss of heat to the surroundings. The process is called calorimetry.

Example

If 100 g of alcohol at 50°C is mixed with 250 g of water at 20°C, what is the final temperature of the mixture?

Solution

$$\begin{aligned}
Q \text{ lost (alcohol)} &= Q \text{ gained (water)} \\
(mc\Delta T)_{\text{alcohol}} &= (mc\Delta T)_{\text{water}} \\
-0.1 \times 2450 \times (T_f - 50) &= 0.250 \times 4200 \times (T_f - 20) \\
-245T_f + 12\,250 &= 1050T_f - 21\,000 \\
12\,250 + 21\,000 &= (1050 + 245)T_f \\
33\,250 &= 1295T_f \\
T_f &= 25.5°\text{C}
\end{aligned}$$

Note: the negative sign indicates a loss.

✔ TEST YOUR UNDERSTANDING

A common mistake among lower primary students is to say that, when two identical glasses of water both at 40°C are mixed, the final temperature will be 80°C. *Write an explanation suitable for a Grade 3 student about why this is not true. Then explain why you add the volumes together (1 cup + 1 cup = 2 cups).*

Questions

10 A 50 g copper mass is heated by placing it in boiling water. It is then placed in a beaker containing 250 g of an unknown liquid at 20°C. The final temperature of the weight and the liquid is found to be 25°C. What is the specific heat of the liquid? (Assume no heat is lost to the surroundings.)

11 A 1500 W electric jug is used to heat 500 mL of water. Calculate the time for the jug to raise the temperature of the water from room temperature, 20°C, to boiling point. (*Note:* 1500 W means it supplies 1500 J of heat energy every second.) The density of water is 1000 kg m^{-3}.

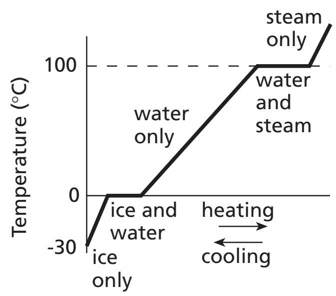
12 In an experiment, 500 g of copper at 80°C is dropped into 1 kg of kerosene at 20°C. The mixture reached a temperature of 25°C. What is the specific heat capacity of kerosene?

10.7 CHANGE OF STATE

Up to now we have only considered substances changing their temperature as heat is added or taken away. The objects remain in the same state of matter. We will now look at what happens when substances change state, that is, change from a solid to a liquid, a liquid to a gas, or vice versa. This is also called change of 'phase'.

What is happening when a block of ice changes to water involves an understanding of thermal energy and the structure of matter. In Section 10.2 the internal or thermal energy of a substance was defined as the total energy possessed by the particles of the substance. This is made up of both kinetic and potential energies. In solids, like a block of ice, the particles are held firmly in position by the bonds between the particles. They contain kinetic energy in the form of vibrational motion, as well as several forms of potential energy. As the ice is heated, the vibrational motion and therefore the kinetic energy and temperature increase. As the particles vibrate faster they spread apart, also increasing their potential energy. As heating continues the vibrations increase until the molecular forces are no longer strong enough to hold the particles together in fixed positions. The particles break free and are able to slide past one another. The solid melts. It requires a large amount of energy to break the bonds and increase the potential energy of the particles. When this is occurring the addition of thermal energy does not go into changing the kinetic energy of the particles but into increasing the potential energy. Since temperature is a measure of the average kinetic energy of the particles, the temperature does not increase. Thus when a solid melts by the addition of heat, the potential energy of the particles increases without a change in temperature. This is shown graphically in Figure 10.15.

Figure 10.15
An effect of heat on the three states of water.



The amount of energy required to melt 1 kg of a substance is called the **specific latent heat of fusion**. (The word 'latent' means hidden.) Why is this heat hidden?

To change 1 kg of ice at 0°C to water at 0°C requires 3.34×10^5 J of energy. Therefore ice has a specific latent heat of fusion of 3.34×10^5 J kg^{-1}.

The reverse is also true; to change 1 kg of water at 0°C to ice at 0°C, that is, to freeze the water, requires the removal of 3.34×10^5 J of energy. The specific latent heat of fusion differs for different substances because of their different bonding. The specific latent heats of fusion of some substances are given in Table 10.2.

Table 10.2 TYPICAL LATENT HEATS

SUBSTANCE	SPECIFIC LATENT HEAT OF FUSION L_f (J kg^{-1})	SPECIFIC LATENT HEAT OF VAPORISATION L_v (J kg^{-1})
Mercury	1.18×10^4	2.90×10^5
Lead	2.30×10^4	8.64×10^5
Gold	6.30×10^4	1.64×10^6
Silver	1.05×10^5	2.36×10^6
Alcohol	1.09×10^5	8.70×10^5
Aluminium	1.80×10^5	1.14×10^7
Copper	2.05×10^5	4.82×10^5
Iron	2.76×10^5	6.29×10^6
Water	3.34×10^5	2.25×10^6

After melting, the addition of heat results in an increase in the kinetic energy (now translational, rotational and vibrational) as well as the potential energies of the liquid. Thus temperature again rises, as shown in Figure 10.15.

As the temperature increases some particles begin to break the cohesion forces holding them together. The forces of attraction between the particles become very weak and the particles move more freely. The substance changes state from a liquid to a gas. At a certain temperature any added thermal energy goes into changing the potential energy of the

NOVEL CHALLENGE

At the Le Mans race in France there is 14 km of track and cars reach 360 km h^{-1} (100 m s^{-1}). At the end of the straight drivers approach Mulsanne Corner at 250 km h^{-1} and jam on their brakes to go through the corner at 56 km h^{-1}. The disc rotors glow red hot and sometimes reach 800°C. If a disc rotor (there are four) is 30 cm in diameter and made from steel 0.8 cm thick (density 7.8 g/cm^3), and is 400°C before the corner, show that 95% of the kinetic energy is transferred to heat.

particles, causing the particles to break the cohesion forces. Again at this point the temperature does not increase as there is no increase in the kinetic energy of the particles. This is shown in Figure 10.15. This temperature is called the **boiling point** of the liquid.

It again requires large amounts of thermal energy to change 1 kg of a liquid to a gas or vapour. For example, it requires 2.26×10^6 J of energy to change 1 kg of water at 100°C into steam at 100°C. The thermal energy required to bring about this change is called the **specific latent heat of vaporisation**.

Specific latent heats of vaporisation of other liquids are given in Table 10.2.

The reverse is again true. It requires the removal of 2.26×10^6 J of thermal energy to change 1 kg of steam into water without change in temperature.

The heat required to melt a mass of a substance is given by the equation:

$$Q = mL_f$$

where Q is the heat required in J, m is the mass of the substance in kg, L_f is the specific latent heat of fusion of the substance in J kg^{-1}.

Similarly, the energy required to vaporise a liquid is given by the equation:

$$Q = mL_v$$

where L_v is the specific latent heat of vaporisation.

Example 1

How much energy is required to change a 2 kg block of lead to liquid at its melting point?

Solution

$$\begin{aligned} Q &= mL_f \\ &= 2 \times 2.30 \times 10^4 \\ &= 4.60 \times 10^4 \text{ J} \end{aligned}$$

Example 2

An ice tray containing 200 g of water at 25°C is placed in the freezer. How much heat energy has to be removed to change the water into ice at –4°C?

Solution

$$\begin{aligned} Q &= (mc\Delta T)_{water} + mL_f + (mc\Delta T)_{ice} \\ &= 0.2 \times 4200 \times (25 - 0) + 0.2 \times 3.35 \times 10^5 + 0.2 \times 2060 \times (4 - 0) \\ &= 21\,000 + 0.668 \times 10^5 + 1648 \\ &= 8.94 \times 10^4 \text{ J} \end{aligned}$$

Questions

13 Find the energy required to melt 2.5 kg of gold at its melting point.

14 Copper has a melting point of 1083°C. Find the energy required to melt 200 g of copper originally at room temperature of 22°C.

15 A 2.0 L bottle of water at 20°C is placed in the freezer of a refrigerator. How much heat must be removed by the refrigerator to freeze this water?

16 A child wanting to make a cordial ice block places 200 g of cordial at 25°C in the freezer. If the freezer can remove energy at the rate of 25 joules per second, what time will it take for the cordial to freeze? (Assume the specific latent heat and specific heat capacity of cordial are the same as water.)

17 Two ice blocks of mass 20 g each are placed in 500 g of water at 40°C. What will be the final temperature of the mixture? (Assume no heat is lost to the container or the surroundings.)

10.8 CHANGING THE MELTING AND BOILING POINTS

Most substances have fixed melting and boiling points as long as they are in pure form. Values given in scientific data are for pure substances. However, the melting point and boiling points can be changed by adding impurities to the substance or by pressure changes. For example, if salt is added to water the melting point is lowered. That is, it freezes at a temperature lower than 0°C.

- Why do councils in some cold countries put salt on the road?
- Why do motorists add another liquid to their radiator in these countries?

Adding impurities raises the boiling point of liquids. If salt is added to water, the water will boil above 100°C.

- How does this affect the cooking rate of potatoes and pasta that are cooked in salty water?

Pressure changes also affect the freezing and boiling points of liquids. Increased pressure lowers the freezing point of liquids that expand when they freeze, such as water, but it raises the freezing point for those that contract when they freeze. This is because increasing pressure pushes the molecules closer together, therefore increasing the temperature at which molecules are attracted to each other and form solids. The following are two examples of this phenomenon.

- If two ice blocks are squeezed together they will melt, but if this pressure is released the water between them will refreeze, gluing them together.
- A thin wire with two weights will cut easily through a block of ice. Why?

Pressure changes also affect the boiling point of a liquid. Increasing the pressure increases the boiling point of a liquid, and decreasing the pressure decreases the boiling point.

- Why is it hard to get a good cup of tea on the top of Mount Everest?

Pressure cookers rely on this property. When the lid is placed on the pressure cooker and the cooker is heated, the pressure can be about twice the normal air pressure. The water boils at about 120°C, and therefore the food cooks faster.

Where else is this principle used?

10.9 EVAPORATION

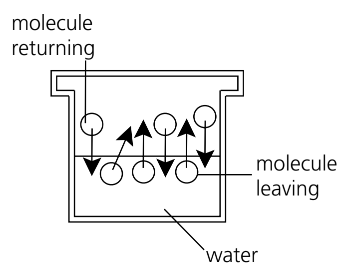
Liquids can change state without boiling. This process is going on all the time in nature. Puddles of water on the road dry up even when the weather is cool. Aftershave and perfume soon disappear from the skin. This is the process of evaporation.

The latent heat of vaporisation plays an important part in evaporation. For molecules of a liquid to change state and become gaseous molecules they require energy. All the molecules in a liquid do not have the same kinetic energy; when molecules evaporate, the faster (hotter) molecules near the surface of the liquid leave first and the slower molecules remain behind. So when water evaporates from the skin your skin feels cooler because the average kinetic energy of those molecules remaining is less. This process, together with sweating, acts as a cooling mechanism for our bodies. When the sweat evaporates from our skin, the skin feels cooler. If there is a breeze blowing, the sweat evaporates faster producing a greater cooling effect.

How can you tell the direction of a breeze by holding up a moist finger?

Evaporation from a container can be stopped or reduced by putting on the lid. In a closed container two processes are occurring: the evaporation of molecules from the liquid and the condensation of molecules back to liquid. If the gaseous molecules are not removed, an equilibrium is reached in which the number of molecules leaving the liquid is equal to the number going back into the liquid. (See Figure 10.16.)

Figure 10.16
In a closed container the rate of evaporation is equal to the rate of condensation, reducing the loss of liquid.



Questions

18 A few years ago it was not uncommon to see cars travelling in country areas with porous water bags made of canvas attached to the front of the car. Why was the temperature of the water in these bags cooler than that of the air?

19 Why does your skin feel cool after placing an alcohol-based aftershave or perfume on it? This is even more evident if you are in a breeze or you blow on it. Is this related to the pain-relief sprays used by footballers?

LAWS OF THERMODYNAMICS 10.10

What happens when you place a cold body in contact with a warm one, for example, placing an ice cube in a glass of water?

The answer to this has already been discussed. As the molecules collide energy is transferred from those that contain the most to those that contain less, that is, from the hot body to the cool one until equilibrium is reached.

Heat flows from the hotter body to the cooler one.

Under perfect conditions where energy is not lost to the surroundings, the energy lost by the hot body is equal to that gained by the cool one.

What are other ways of 'heating' up a cool body, that is, giving the molecules of the cool body more energy? Think of how you heat up your hands on a cold morning. By rubbing your hands together the mechanical energy of your moving hands, with friction, causes your hands to become hotter.

James Joule (as previously discussed) increased the temperature of water by using the potential energy of the falling weight.

The thermal energy of a system can be increased by adding heat to it or by doing work on it.

The total increase in the thermal energy of an isolated system is equal to the sum of the heat added to it and the work done on it. This is the **first law of thermodynamics**. Notice this is just an extension of the principle of conservation of energy.

The **second law of thermodynamics** formulated by a German physicist Rudolf Clausius (1822–88) relates heat transfer to differences in temperature. For example, what are some devices that take heat out of objects or the air?

Some you may have guessed are: refrigerators, airconditioners, freezers, or even the cooling system of a car.

Each of these devices takes heat from something and transfers it to something of lower temperature. What would happen if the car was driven in a place where the atmospheric temperature was higher than the temperature of the car engine?

This is one of the reasons it is suggested that it is impossible to obtain a temperature of zero kelvins. There is no place to transfer the heat, no point of lower temperature.

To transfer internal energy from a low temperature heat source to a higher one requires work, such as the work done by the motor in a refrigerator or an air conditioner.

Consider now a hot body which is placed in contact with a cold body. Heat is transferred from this hot body to the cold body and after a time equilibrium is reached. If it is an isolated system there is no loss in energy in the process but the system has lost its capacity to do work. There is no longer as great a temperature difference. This is seen in the example of placing a block of ice in hot water. Before placing the ice in the water the molecules of the ice are in a well-ordered crystalline arrangement. After the equilibrium temperature is reached the molecules are in a less ordered, more random motion. The molecules of the ice have become less ordered. The scientific term that defines the orderliness of the molecules is **entropy**.

Entropy is a measure of the disorder of a system. The more disorder the greater the entropy. Natural processes always go in a direction that causes an increase in the total

INVESTIGATING

How could you measure the uneven heating across the carousel inside a microwave if you were given a thermometer, some water and a dozen little plastic containers that rolls of film come in? Be extremely careful if you choose to try it. Start with a low time first, say 20 seconds. Why?

INVESTIGATING

After inventing the first, second and third laws of thermodynamics, physicists decided they had a 'zeroth law'. What on Earth is that?

✔ TEST YOUR UNDERSTANDING

(Answer true or false)

- Heat and temperature are the same thing.
- Heat and cold flow like liquids.
- The hotter of two objects contains the more heat.

entropy of the universe. This is the second law of thermodynamics. This law would indicate that the availability of energy in the universe is decreasing.

You can, of course, have a decrease in entropy in one part of the universe (e.g. freezing water) but there is an overall greater increase in entropy elsewhere (e.g. heat produced at the rear of the refrigerator).

— Practice questions

The relative difficulty of these questions is indicated by the number of stars beside each question number: * = low; ** = medium; *** = high.

Review — applying principles and problem solving

***20** Compare the temperature and thermal energy of a cup of coffee at 90°C and a swimming pool at 22°C.

***21** Four different masses of the same metal, 100 g, 200 g, 500 g, and 1 kg, were heated by a 2000 J energy supply. If the 100 g mass changed its temperature by 5°C, what would have been the temperature changes of the other three?

***22** State two reasons why a mercury-in-glass thermometer could not be used to measure the temperature of a pottery kiln.

***23** Convert the following Celsius temperatures to kelvins: **(a)** 290°C; **(b)** –25°C; **(c)** 59.2°C.

***24** Change the following Kelvin temperatures to °C: **(a)** 69 K; **(b)** 1376 K; **(c)** 345.6 K.

***25** 1500 J of energy is used to heat a 400 g sample of iron initially at 28°C. What would be the final temperature of the iron?

****26** A beaker containing 200 g of mercury at 15°C was placed in a freezer. Find the energy removed from the mercury if it cooled down to –4°C. Would the freezer have to work harder if the beaker contained 200 g of water instead of mercury? Explain!

****27** An electric kettle was used to heat 500 g of water at 20°C. If the kettle can supply energy at the rate of 1500 J per second, and was turned on for one and a half minutes, what was the final temperature of the water?

****28** A 200 g bar of aluminium was heated in a bunsen burner until its temperature was 150°C. It was then plunged into a beaker containing 500 g of paraffin at 50°C. What was the final temperature of the mixture?

***29** Why is heat needed to change a solid to a liquid but the substance does not change temperature?

***30** Explain why steam at 100°C would be more effective in heating a cup of milk to make coffee (a cup of cappuccino, for example) than water at 100°C?

****31** Dry steam is used to make a cup of coffee by bubbling it through water. If the steam is at 100°C, what mass of steam must be used to heat 200 g of water from 25°C to 95°C?

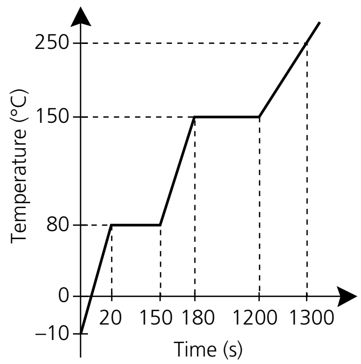
****32** An electric kettle, rated at 2 kW, is filled with water and its total mass determined. The kettle is switched on and the water is allowed to boil for a further 60 s after coming to the boil. The kettle is found to be 80 g lighter. Calculate the specific latent heat of vaporisation. Suggest why this value differs from the stated value of the latent heat of vaporisation.

***33** Suggest the weather conditions that would make clothes on an outside line dry faster.

***34** State two ways in which boiling is different from evaporation.

***35** State two ways in which the rate of evaporation can be increased.

Figure 10.17
A heating curve for question 36.



36 Figure 10.17 shows the relationship between the temperature of 100 g of an unknown substance and the time it is heated by a 1000 W source.
(a) What is the melting point of the substance?
(b) What is the boiling point of the substance?
(c) How much energy is required to melt the substance?
(d) What is the specific latent heat of fusion of this substance?
(e) What is the specific latent heat of vaporisation of this substance?
(f) Calculate the specific heat capacity of the substance in the liquid state.

****37** While carrying out an experiment to measure the boiling point of water at various altitudes, students found that instead of the boiling point decreasing as they went higher in the mountains, the temperature rose slightly. Suggest why this might have occurred.

****38** Students performing an experiment on naphthalene to discover how its temperature changed with time allowed hot naphthalene to cool down. Table 10.3 lists their results.

Table 10.3

Time (minutes)	0	1	2	3	4	5	6	7	8	9	10
Temperature (°C)	103	93	83	80	80	80	80	75	70	65	60

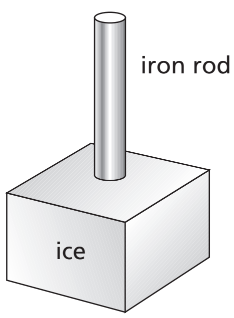
(a) Plot the graph of temperature against time.
(b) Explain what is happening between the times of 3 minutes and 6 minutes.

Extension — complex, challenging and novel

*****39** An electric shower unit is rated at 5 kW. If water enters it at 15°C and leaves it as hot water at the rate of 5 kg per minute, what is the temperature of the hot water?

*****40** An iron rod of mass 100 g and 4 cm in diameter was placed in a furnace and heated until its temperature reached 150°C. It was then placed on its end on top of a large block of ice (temperature of 0°C). (See Figure 10.18.) How far into the block of ice will the rod sink (assume no heat is lost to the surroundings).

Figure 10.18
For question 40.



*****41** A block of ice of temperature 0°C and mass 20 g was placed in a beaker and weighed. The total mass was 55 g. Steam at 110°C was ducted onto the ice until the ice completely melted. Assuming no loss of heat to the surroundings, find the mass of the beaker and its contents!