



Energy released by the combustion of fuels is used for everyday activities like heating, cooking, and the generation of electricity. Humans also consume food each day to obtain fuel for movement and all the functions of the body.

While we depend on the energy obtained from fuels, their combustion also releases gases such as carbon dioxide and water vapour which are both greenhouse gases. An increase of these in the atmosphere has serious environmental implications.

In this chapter, you will learn about the energy content in fuels and food. You will learn how to calculate the amount of useful energy produced when a certain mass of a fuel or food undergoes combustion and you will also learn how to calculate the mass and volume of products such as the major greenhouse gases.

### Key knowledge

- determination of limiting reactants or reagents in chemical reactions **3.2**
- calculations related to the application of stoichiometry to reactions involving the combustion of fuels, including mass–mass, mass–volume and volume–volume stoichiometry, to determine heat energy released, reactant and product amounts and net volume or mass of major greenhouse gases ( $\text{CO}_2$ ,  $\text{CH}_4$  and  $\text{H}_2\text{O}$ ), limited to standard laboratory conditions (SLC) at  $25^\circ\text{C}$  and 100 kPa **3.1**
- the use of specific heat capacity of water to approximate the quantity of heat energy released during the combustion of a known mass of fuel and food **3.3**
- the principles of solution calorimetry, including determination of calibration factor and consideration of the effects of heat loss; analysis of temperature–time graphs obtained from solution calorimetry **3.4**
- energy from fuels and food:
  - calculation of energy transformation efficiency during combustion as a percentage of chemical energy converted to useful energy
  - comparison and calculations of energy values of foods containing carbohydrates, proteins and fats and oils. **3.5**

VCE Chemistry Study Design extracts © VCAA (2022); reproduced by permission.



## 3.1 Stoichiometry involving combustion of fuels

The ability to measure and predict quantities involved in chemical reactions is a very important part of chemistry. One way scientists evaluate fuels is through comparison of quantities of chemicals involved in combustion reactions, as well as quantities of energy released. For example, the ethanol produced from corn in the plant shown in Figure 3.1.1 can be compared to petroleum through calculations of how much oxygen is required for **complete combustion**, the amount of atmospheric pollution the fuels will produce and the heat energy released per gram of fuel. In this section you will learn about these types of calculations, all of which rely on an understanding of the mole concept.



**FIGURE 3.1.1** A corn ethanol processing plant. Ethanol is a biofuel that is used as an alternative to petroleum. The concepts taught in this section allow you to calculate the mass of oxygen required to combust a given mass of ethanol, the mass of carbon dioxide (a greenhouse gas) produced, as well as the energy released.

### STOICHIOMETRY AND QUANTITIES OF CHEMICALS

The calculations involved in determining the quantity of a reactant or product from another reactant or product use **stoichiometry**. Stoichiometry is the study of mole ratios of substances and is based on the law of conservation of mass.

A chemical reaction involves the rearrangement of pre-existing atoms; atoms are neither created nor destroyed. This means the total mass of all products is equal to the total mass of all reactants.

Consider the equation for the reaction that occurs when methane ( $\text{CH}_4$ ) burns in oxygen:



The **coefficients** used to balance the equations show the **mole ratio** between the reactants and products involved in the reaction. The equation indicates that one mole of  $\text{CH}_4(\text{g})$  reacts with two moles of  $\text{O}_2(\text{g})$  to form one mole of  $\text{CO}_2(\text{g})$  and two moles of  $\text{H}_2\text{O}(\text{g})$ . Examples of what this means in more general terms are:

- the amount of oxygen required to react with the methane will be double the amount of methane used
- the amount of carbon dioxide produced will be the same as the amount of methane that reacted.

These mole ratios can be expressed in formulas:

$$\frac{n(\text{O}_2)}{n(\text{CH}_4)} = \frac{2}{1} \text{ and } \frac{n(\text{CO}_2)}{n(\text{CH}_4)} = \frac{1}{1}$$



## MOLE THEORY

### The mole

The mole is a unit used by chemists for counting particles. One mole contains exactly  $6.022\,140\,772 \times 10^{23}$  particles. We usually round this off to  $6.02 \times 10^{23}$  particles. This number is called Avogadro's constant and is given the symbol  $N_A$ .

The symbol for **amount** of substance measured in moles is  $n$ . The unit is mol.

A useful relationship links the amount of a substance ( $n$ ), in mol, Avogadro's constant ( $N_A$ ) and the number of particles in a substance ( $N$ ):

$$n = \frac{N}{N_A}$$

Where needed, this relationship can be rearranged to:

$$N = n \times N_A$$

### Mass

Chemists often use mass, measured in grams, to measure an exact amount, in mol, of an element or compound.

The molar mass of an element or compound is the mass of one mole of the element or compound. The amount of a substance ( $n$ ), in mol, is related to the mass of the substance ( $m$ ), in g, and its molar mass ( $M$ ) in  $\text{g mol}^{-1}$ , by the formula:

$$n = \frac{m}{M}$$

This relationship can be rearranged to:

$$m = n \times M \text{ or } M = \frac{m}{n}$$

### Volume of gases

Quantities of gases are often measured using volume, measured in litres. The same amount, in mol, of different types of gases will occupy the same volume if pressure and temperature are the same. The **molar volume** of a gas is the volume of one mole of the gas. At standard laboratory conditions (SLC) of 100 kPa and 25°C, the molar volume has been determined to be  $24.8 \text{ L mol}^{-1}$ .

A useful relationship links the amount of a gas ( $n$ ), in mol, its molar volume ( $V_m$ ) in  $\text{L mol}^{-1}$ , and the volume occupied by the gas ( $V$ ) in L:

$$n = \frac{V}{V_m}$$

Where needed, this relationship can be rearranged to:

$$V = n \times V_m$$

$24.8 \text{ L mol}^{-1}$  can be substituted into either formula for  $V_m$  if the gas is at SLC.

In general, for stoichiometric calculations you will be given, or will be able to work out, the number of moles of one chemical in the reaction (called the 'known chemical'). This can then be used with the mole ratio to determine the number of moles of one of the other reactants or products involved in the reaction (called the 'unknown chemical').

The mole ratio can be written:

$$\frac{n(\text{unknown chemical})}{n(\text{known chemical})} = \frac{\text{coefficient of unknown chemical}}{\text{coefficient of known chemical}}$$

## MASS-MASS STOICHIOMETRY

When a reaction is carried out in a laboratory, quantities of chemicals are often measured in grams, not moles. As stoichiometry uses a mole ratio, the amount, in mol, of the known chemical must first be calculated from its mass using the formula:

$$n = \frac{m}{M}$$

After the mole ratio has been used to determine the amount, in mol, of the unknown chemical, the mass of the unknown chemical is calculated by rearranging the formula to:

$$m = n \times M$$

When the unknown and the known in a stoichiometric calculation are both masses, the calculation is known as **mass-mass stoichiometry**.

**i** The coefficients used to balance the equations show the mole ratio between the reactants and products involved in the reaction.

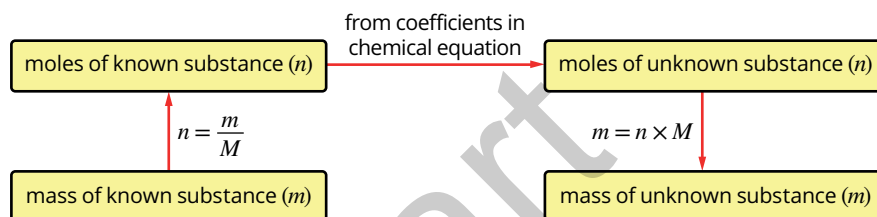
## Calculating the mass of carbon dioxide produced in a combustion reaction

Stoichiometry can be combined with your knowledge of combustion reactions to find the mass of carbon dioxide produced.

There are several steps involved in calculating the mass of carbon dioxide produced based on the mass of fuel that undergoes combustion. The mass of fuel is the known chemical and carbon dioxide is the unknown.

- 1 Write a balanced chemical equation for the reaction.
- 2 Calculate the number of moles of the fuel from its mass.
- 3 Find the mole ratio from the coefficients in the chemical equation to calculate the number of moles of carbon dioxide produced.
- 4 Calculate the mass of carbon dioxide produced.

Figure 3.1.2 provides a flow chart that summarises this process. Worked example 3.1.1 will help you understand these steps.



**FIGURE 3.1.2** A flow chart for mass–mass stoichiometric calculations is helpful when solving these types of problems.

### Worked example 3.1.1

#### MASS–MASS STOICHIOMETRIC PROBLEMS

Calculate the mass of carbon dioxide, in kg, produced when 540 g of propane ( $C_3H_8$ ) burns completely in oxygen.

Thinking	Working
Write a balanced equation for the reaction.	$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$
As the quantity of the known substance is given in g, calculate the number of moles of the known substance using: $n = \frac{m}{M}$	$n(C_3H_8) = \frac{540}{44.0}$ $= 12.3 \text{ mol}$
Find the mole ratio: $\frac{\text{coefficient of unknown}}{\text{coefficient of known}}$	$\frac{n(CO_2)}{n(C_3H_8)} = \frac{3}{1}$
Calculate the number of moles of the unknown substance using: $n(\text{unknown}) = \text{mole ratio} \times n(\text{known})$	$n(CO_2) = \frac{3}{1} \times 12.3$ $= 36.8 \text{ mol}$
Calculate the mass of the unknown substance using: $m = n \times M$	$m(CO_2) = 36.8 \times 44.0$ $= 1620 \text{ g}$ $= 1.62 \text{ kg}$

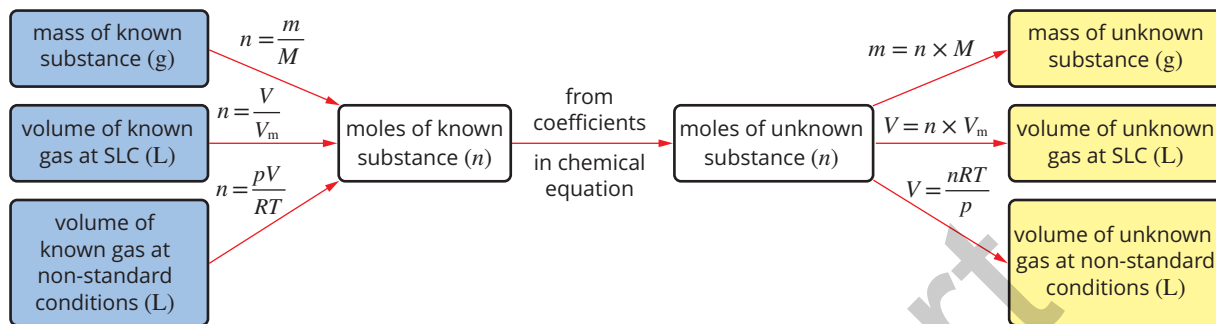
### Worked example: Try yourself 3.1.1

#### SOLVING MASS–MASS STOICHIOMETRIC PROBLEMS

Calculate the mass of carbon dioxide, in kg, produced when 3.60 kg of butane ( $C_4H_{10}$ ) burns completely in oxygen.

## MASS-VOLUME STOICHIOMETRY

Stoichiometric calculations that follow the same general pattern can also be used to calculate the volume of oxygen required to combust a fuel and the volumes of gases produced by the reactions. When the known in a stoichiometric calculation is a mass and the unknown is a volume of a gas, the calculation is known as **mass-volume stoichiometry**. The number of moles of the known substance (often mass of fuel) is calculated from data that is given to you, the mole ratio from the coefficients in the equation is used to find the number of moles of the unknown chemical, and the desired quantity of the unknown substance is then calculated. This is summarised in Figure 3.1.3.



**FIGURE 3.1.3** Stoichiometric calculations generally follow the steps shown in the flow chart. Calculating the number of moles and using a mole ratio from a balanced chemical equation are central to all stoichiometric calculations.

Worked example 3.1.2 shows how to calculate the volume of carbon dioxide under **standard laboratory conditions, SLC**, produced from a known mass of fuel. Standard laboratory conditions for a gas indicate that the gas has a pressure of 100 kPa and that the temperature is 298 K (100°C).

### Worked example 3.1.2

#### MASS-VOLUME STOICHIOMETRIC CALCULATIONS AT SLC

Calculate the volume of carbon dioxide, in L, produced when 2.00 kg of propane ( $C_3H_8$ ) is burned completely in oxygen. The gas volume is measured at SLC.

Thinking	Working
Write a balanced chemical equation for the reaction.	$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(l)$
As the quantity of the known substance is a mass, calculate the number of moles of the known substance using: $n = \frac{m}{M}$	$n(C_3H_8) = \frac{2000}{44.0}$ $= 45.5 \text{ mol}$
Find the mole ratio: $\frac{\text{coefficient of unknown}}{\text{coefficient of known}}$	$\frac{n(CO_2)}{n(C_3H_8)} = \frac{3}{1}$
Calculate the number of moles of the unknown substance using: $n(\text{unknown}) = \text{mole ratio} \times n(\text{known})$	$n(CO_2) = \frac{3}{1} \times 45.5$ $= 136 \text{ mol}$
Calculate the volume of the unknown substance using: $V = n \times V_m$	$V(CO_2) = 136 \times 24.8$ $= 3382 \text{ L}$ $= 3.38 \times 10^3 \text{ L}$

### Worked example 3.1.2: Try yourself

#### MASS-VOLUME STOICHIOMETRIC CALCULATIONS AT SLC

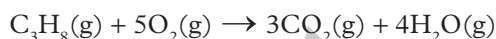
Calculate the volume of carbon dioxide, in L, produced when 300 g of butane ( $C_4H_{10}$ ) is burned completely in oxygen. The gas volume is measured at SLC.

### GAS VOLUME-VOLUME CALCULATIONS

For chemical reactions where all reactants and products are gases, it is often more convenient to measure reactant volumes, rather than masses. When the unknown and the known in a stoichiometric calculation are both gas volumes, the calculation is known as **volume-volume stoichiometry**.

All gases occupy equal volumes measured at the same temperature and pressure. Therefore, the mole ratios of gases will also be volume ratios if temperature and pressure are kept constant.

For example, the reaction between propane gas and oxygen can be represented by the equation:



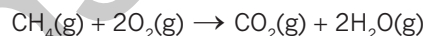
This equation tells us that one mole of propane reacts with five moles of oxygen gas to produce three moles of carbon dioxide and four moles of water vapour. As every reactant and product is in the gaseous state, the ratio also tells us that one litre of propane gas reacts with five litres of oxygen gas to produce three litres of carbon dioxide and four litres of water vapour at constant temperature and pressure.

**i** When all species in a chemical equation are in the gaseous state, and the temperature and pressure are constant, the mole ratio will also be a volume ratio.

### Worked example 3.1.3

#### GAS VOLUME-VOLUME CALCULATIONS

Methane gas ( $CH_4$ ) is burned in a gas stove according to the following equation:



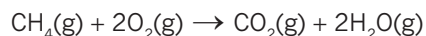
If 50 mL of methane is burned, calculate the volume of  $O_2$  gas required for complete combustion of the methane under constant temperature and pressure conditions.

Thinking	Working
Find the mole ratio: $\frac{\text{coefficient of unknown}}{\text{coefficient of known}}$	$\frac{n(O_2)}{n(CH_4)} = \frac{2}{1}$
The temperature and pressure are constant, so volume ratios are the same as mole ratios.	$\frac{V(O_2)}{V(CH_4)} = \frac{2}{1}$
Calculate the volume of the unknown substance using: $V(\text{unknown}) = \text{mole ratio} \times V(\text{known})$	$\begin{aligned} V(O_2) &= \frac{2}{1} \times 50 \\ &= 100 \text{ mL} \end{aligned}$

### Worked example: Try yourself 3.1.3

#### GAS VOLUME-VOLUME CALCULATIONS

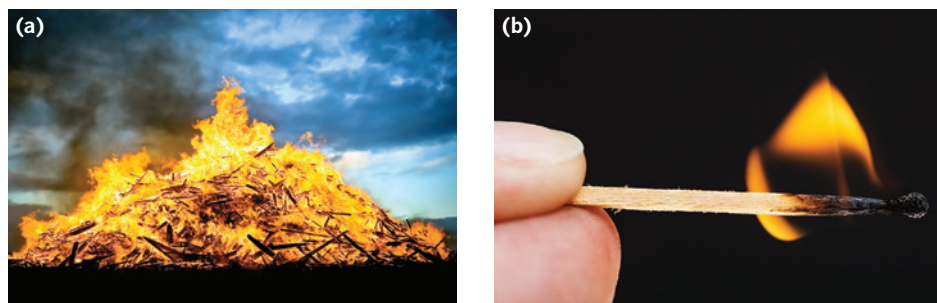
Methane gas ( $CH_4$ ) is burned in a gas stove according to the following equation:



If 50 mL of methane is burned in air, calculate the volume of  $CO_2$  gas produced under constant temperature and pressure conditions.

## STOICHIOMETRY AND QUANTITY OF ENERGY RELEASED

The quantity of energy released when a fuel undergoes combustion is directly proportional to the amount of fuel used. In Figure 3.1.4, the bonfire contains more fuel than the match, so the bonfire releases far more energy. We also know that doubling the quantity of petrol in a car allows a car to travel twice as far.



**FIGURE 3.1.4** The amount of energy released by a bonfire is far greater than the amount of energy released by the burning of a single match.

We can use a thermochemical equation and the principles of stoichiometry to calculate the energy released by the combustion of specified quantities of fuel. The use of a thermochemical equation means energy released can be calculated and compared under different conditions.

### Combustion and energy

The quantity of energy obtained from the combustion of a fuel depends on:

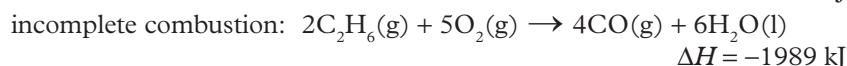
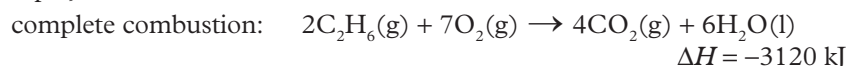
- the type of fuel
- the amount of fuel
- whether complete combustion or incomplete combustion is involved.

The thermochemical equations for the complete combustion of methane and pentane are shown below. Complete combustion occurs when there is a plentiful oxygen supply.



The equations show that the equal amounts of different fuels release different quantities of energy. The combustion of one mole of pentane releases more energy than the combustion of one mole of methane.

**Incomplete combustion** occurs when the oxygen supply is limited, and carbon monoxide is formed rather than carbon dioxide. The complete combustion of a fuel releases more energy than incomplete combustion of the same fuel. For example, for ethane:



### Calculating energy change from thermochemical equations

The coefficients of the reactants in a thermochemical equation indicate the amount, in moles, of each substance that reacts or is produced to give the specified enthalpy change,  $\Delta H$ .

For example, the thermochemical equation for the complete combustion of ethane indicates that 3120 kJ of energy is released by the reaction of 2 moles of ethane with 7 moles of oxygen.

Worked examples 3.1.4 and 3.1.5 on the following pages show you how to use a thermochemical equation to calculate the energy released by the combustion of different amounts of fuels or the amount of fuel required to produce a specified quantity of energy.

**i** Complete combustion produces carbon dioxide and water as products. Incomplete combustion forms carbon monoxide and sometimes carbon.

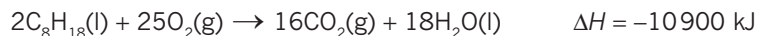
**i** The coefficients of the reactants in a thermochemical equation indicate the amount, in moles, of each substance that react to give the enthalpy change,  $\Delta H$ , specified in the equation.



### Worked example 3.1.4

#### USING A THERMOCHEMICAL EQUATION TO CALCULATE ENERGY RELEASED BY THE COMBUSTION OF A FUEL

Calculate the heat energy released, in MJ, when 10.0 kg of octane undergoes complete combustion.

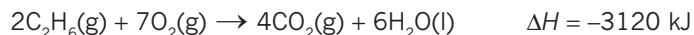


Thinking	Working
As the quantity of the known substance is a mass, calculate the number of moles of the known substance using: $n = \frac{m}{M}$	$n(\text{C}_8\text{H}_{18}) = \frac{10.0 \times 10^3}{114.0}$ $= 87.7 \text{ mol}$
Using the thermochemical equation, find the relationship between the number of moles of fuel burned and energy released.	x kJ is released by 87.7 mol 10900 kJ is released by 2 mol
Calculate the energy released by the fuel in kilojoules.	By proportion: $\frac{x}{10900} = \frac{87.7}{2}$ $x = \frac{87.7}{2} \times 10900$ $= 477\,965 \text{ kJ}$
Convert the energy released to MJ using: $1 \text{ MJ} = 10^3 \text{ kJ}$ Express the answer to the correct number of significant figures. 10.0 kg of octane in question has 3 significant figures. Therefore, 3 significant figures are given in the answer.	477 965 kJ = 478 MJ (to 3 significant figures)

### Worked example: Try yourself 3.1.4

#### USING A THERMOCHEMICAL EQUATION TO CALCULATE ENERGY RELEASED BY THE COMBUSTION OF A FUEL

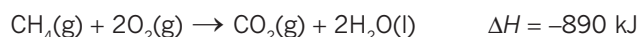
Calculate the heat energy released, in MJ, when 10.0 kg of ethane undergoes complete combustion.



### Worked example 3.1.5

#### USING A THERMOCHEMICAL EQUATION TO CALCULATE THE AMOUNT OF FUEL THAT MUST BE BURNED TO PRODUCE A PARTICULAR AMOUNT OF ENERGY

What volume of methane, measured at SLC, burns completely to provide  $4.00 \times 10^4$  kJ?

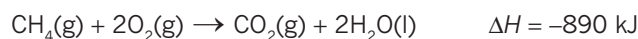


Thinking	Working
Using the thermochemical equation, write a relationship between the number of moles of fuel burned and energy released.	x mol releases $4.00 \times 10^4$ kJ 1 mol releases 890 kJ
Calculate the amount of the fuel that was burned to produce the energy.	By proportion: $\frac{x}{1} = \frac{4.00 \times 10^4}{890}$ $= \frac{4.00 \times 10^4}{890} \times 1$ $= 44.9$ mol
Calculate the volume of the fuel, at SLC.	$V(\text{CH}_4) = n \times V_m$ $= 44.9 \times 24.8$ $= 1115$ L $= 1.11 \times 10^3$ L

### Worked example: Try yourself 3.1.5

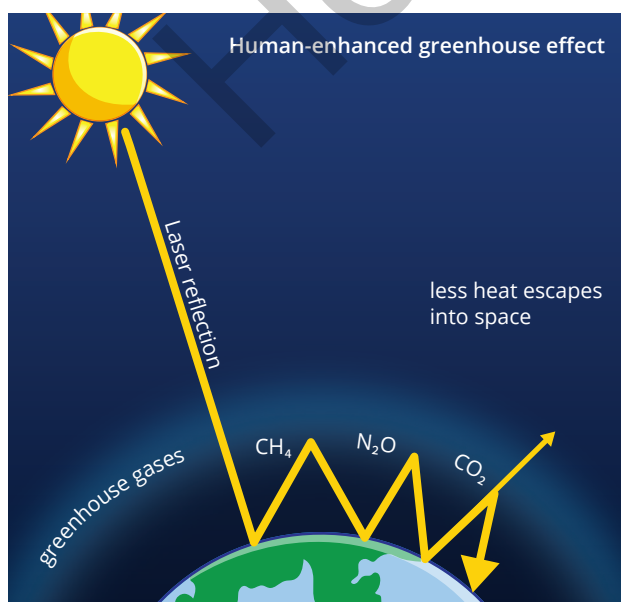
#### USING A THERMOCHEMICAL EQUATION TO CALCULATE THE AMOUNT OF FUEL THAT MUST BE BURNED TO PRODUCE A PARTICULAR AMOUNT OF ENERGY

What volume of methane, measured at SLC, burns completely to provide  $5.00 \times 10^3$  kJ?



## COMBUSTION AND GREENHOUSE GASES

All carbon-based fuels undergo combustion to produce the greenhouse gases carbon dioxide and water. A **greenhouse gas** is a gas that can absorb infrared radiation. Human activities such as combustion of fuels are increasing the presence of greenhouse gases in the atmosphere, which is contributing to an enhanced greenhouse effect that causes global warming. As the amount of greenhouse gases in the atmosphere increases, more heat is trapped closer to Earth, as seen in Figure 3.1.5.



**FIGURE 3.1.5** Human activities are increasing the amount of greenhouse gases in the atmosphere, leading to an enhanced greenhouse effect.

## CHEMFILE

### Energy production in Australia

Many Australian power stations, like the one in the figure below, produce energy from the combustion of brown and black coal. These combustion reactions produce tonnes of carbon dioxide on a daily basis, which is released into the atmosphere. Recent global climate change negotiations among developed nations, including Australia, focus on pledges to reduce the amount of carbon dioxide emitted during the production of energy due to carbon dioxide being a significant greenhouse gas. Stoichiometry allows chemists to usefully predict and compare the mass and volume of carbon dioxide emissions produced by similar quantities of different fuels.

In recent years, a number of Australian states have accelerated plans to close coal-fired power stations due to pressure from cleaner and lower-cost energy generation, including renewable options like solar, wind and batteries. Although these are economic pressures, they have the added benefit of reducing gases released to the atmosphere.



It is useful to be able to compare the amount of carbon dioxide emissions produced during the production of energy from the combustion of different fuels.

For the majority of fuels, the quantities of carbon dioxide and water produced represent the net mass or volume of greenhouse gases added to the atmosphere when those fuels undergo combustion.

Methane, CH<sub>4</sub>, a fuel which is a constituent of both natural gas and biogas, is also a greenhouse gas. In fact, methane is far more potent than both carbon dioxide and water, having a significantly greater warming power than carbon dioxide over the first 20 years after it reaches the atmosphere. Recent observations of methane levels from satellites have shown substantial quantities of methane escaping from landfills and coal mines around the world as well as from agriculture and even home kitchens.

Capturing the methane produced by agricultural animals or decomposition of organic matter in landfill and using it as a biogas in combustion reactions does increase the net volume of carbon dioxide released to the atmosphere, but it also reduces the volume of the more potent methane. Worked example 3.1.6 shows you how to calculate the net volume of greenhouse gases released when methane undergoes combustion.

**i** CO<sub>2</sub>, CH<sub>4</sub> and H<sub>2</sub>O are major greenhouse gases.

### Worked example 3.1.6

#### CALCULATING THE VOLUME OF GREENHOUSE GASES RELEASED

What is the volume of greenhouse gases released when 1.5 L of methane (CH<sub>4</sub>) undergoes complete combustion in oxygen? Assume all reactants and products are in the gaseous state, and the temperature and pressure remain constant.

Thinking	Working
Write a balanced chemical equation for the reaction	$\text{CH}_4(\text{g}) + 2\text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{g})$
Identify the volume ratio between methane and greenhouse gases produced (CO <sub>2</sub> and H <sub>2</sub> O).	$\frac{V(\text{greenhouse gases})}{V(\text{CH}_4)} = \frac{3}{1}$
Calculate the volume of greenhouse gases produced	$V(\text{greenhouse gases}) = \frac{3}{1} \times V(\text{CH}_4)$ $= \frac{3}{1} \times 1.5$ $= 4.5 \text{ L}$
Establish the final volumes of greenhouse gases released.	$V(\text{CH}_4) = 0 \text{ L}$ $V(\text{CO}_2 + \text{H}_2\text{O}) = 4.5 \text{ L}$

### Worked example: Try yourself 3.1.6

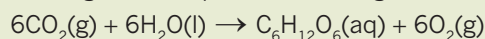
#### CALCULATING THE VOLUME OF GREENHOUSE GASES

What is the volume of greenhouse gases released when 10.8 L of methane (CH<sub>4</sub>) undergoes complete combustion in oxygen? Assume all reactants and products are in the gaseous state, and the temperature and pressure remain constant.



## Bioethanol as a fuel for cars to reduce greenhouse gases

Ethanol,  $C_2H_5OH$ , can be used as a fuel in cars as an alternative to octane (petrol) or as part of a blend with petrol. As with octane,  $C_8H_{18}$ , ethanol produces the greenhouse gases carbon dioxide and water when it undergoes complete combustion. However, when the ethanol is obtained from plant sources it is known as bioethanol and its use has greatly reduced net greenhouse gas emissions. As you learned in Chapter 2, bioethanol is produced from the fermentation of glucose derived from cropping plants such as sugar cane or corn. As the crops grow, they remove carbon dioxide from the atmosphere for the process of photosynthesis, the chemical reaction that produces glucose in plants according to the equation:



This removal of carbon dioxide from the atmosphere offsets the release of carbon dioxide when the ethanol is burnt in cars. The net greenhouse emissions are lower than when petrol obtained from fossil fuels is burnt.

The largest bioethanol producer in Australia is in Nowra on the NSW South Coast (Figure 3.1.6). Australia has the capacity to make approximately 440 million litres of ethanol annually, with nearly 70% of this manufactured at the Nowra facility. The manufacturing process uses a distillery that produces ethanol through the fermentation

of glucose sourced from starch. At the Nowra facility, starch is a by-product of the production of flour from wheat, which is the main manufacturing process at the Nowra site.

The enthalpy of combustion of ethanol is  $-1360 \text{ kJ mol}^{-1}$  at SLC.

### Analysis

- Write a balanced thermochemical equation that represents the complete combustion of ethanol.
- What mass of carbon dioxide, in kg, would be formed if 1.00 kg of ethanol reacts?
  - What volume of carbon dioxide would be formed from 1.00 kg of ethanol at SLC?
- If the density of ethanol is  $0.785 \text{ g mL}^{-1}$ , calculate the mass of ethanol, in kg, in a 50.0 L tank of the fuel using the formula:
 
$$\text{density} = \frac{\text{mass (g)}}{\text{volume (mL)}}$$
  - Calculate the energy, in MJ, that can be obtained from the complete combustion of 50.0 L of ethanol.
- Calculate the mass, in g, of carbon dioxide formed when 1.00 kJ of energy is released.



FIGURE 3.1.6 An ethanol plant at Nowra in New South Wales

## 3.1 Review



### SUMMARY

- The coefficients in a balanced equation show the mole ratio of the amount, in moles, of reactants and products involved in the reaction.
- Stoichiometric calculations follow the general steps:
  - Calculate the amount, in moles, of a known substance from the data given.  
Use  $n = \frac{m}{M}$  or  $n = \frac{V}{V_m}$
  - Use the mole ratio from a balanced chemical equation to determine the amount, in moles, of the unknown substance.  
$$\frac{n(\text{unknown chemical})}{n(\text{known chemical})} = \frac{\text{coefficient of unknown chemical}}{\text{coefficient of known chemical}}$$
  - Find the desired quantity of the unknown substance from its amount, in moles, using  $m = n \times M$  or  $V = n \times V_m$ .
- Stoichiometric calculations can be used to calculate the mass of carbon dioxide and water released during the combustion of a carbon-based fuel and the mass of oxygen required for complete combustion.
- The mole ratio in a balanced equation is also a volume ratio if all reactants and products are in the gaseous state and the temperature and pressure are kept constant.
- Stoichiometric calculations based on thermochemical equations can be used to calculate the amount of heat energy released during the combustion of a carbon-based fuel.
- The coefficients of the reactants in a thermochemical equation indicate the amounts, in moles, of each substance that react to give the enthalpy change specified in the equation.
- Fuels can be compared in terms of energy released and total greenhouse gases produced.

### KEY QUESTIONS

#### Knowledge and understanding

- Consider the following balanced equation that shows the formation of carbon dioxide from carbon monoxide.  
$$2\text{CO}(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{CO}_2(\text{g})$$
  - Calculate the amount, in mol, of oxygen required to completely react with 2 mol of carbon monoxide.
  - Calculate the amount, in mol, of oxygen required to completely react with 6 mol of carbon monoxide.
  - Calculate the volume of oxygen needed to completely react with 1.5 L of carbon monoxide. Assume all volumes are measured at the same temperature and pressure.
- Consider the following thermochemical equation that shows the combustion of methane.  
$$\text{CH}_4(\text{g}) + 2\text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l}) \quad \Delta H = -890 \text{ kJ}$$
  - Determine the amount, in mol, of methane that must undergo combustion in order to release 890 kJ of energy.
  - Calculate the amount, in mol, of carbon dioxide and water that are produced when 890 kJ of energy is released.
  - Calculate the net mass of greenhouse gases released when 890 kJ of energy is released.
- Calculate the energy released when the following quantities of ethane gas burn according to the equation:  
$$2\text{C}_2\text{H}_6(\text{g}) + 7\text{O}_2(\text{g}) \rightarrow 4\text{CO}_2(\text{g}) + 6\text{H}_2\text{O}(\text{l}) \quad \Delta H = -3120 \text{ kJ}$$
  - 3.00 mol
  - 100 g
  - 10.0 L at SLC

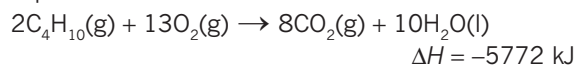
#### Analysis

- Octane ( $\text{C}_8\text{H}_{18}$ ) is a component of petrol. 200 g of octane burns in oxygen to produce carbon dioxide and water. The equation for this reaction is:  
$$2\text{C}_8\text{H}_{18}(\text{g}) + 25\text{O}_2(\text{g}) \rightarrow 16\text{CO}_2(\text{g}) + 18\text{H}_2\text{O}(\text{g})$$
  - Calculate the mass of oxygen required to completely react with 200 g of octane.
  - Calculate the mass of carbon dioxide produced when 200 g of octane is burnt.
- 8.00 g of propane was burned in excess oxygen.
  - Write a balanced chemical equation for the complete combustion of propane.
  - What mass of carbon dioxide would be produced?
  - What volume, at SLC, of oxygen would be consumed in the reaction?

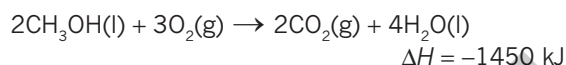
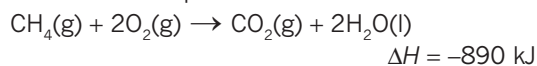
- 6 Calculate the energy released, in MJ, when 250.0 g of petrol burns completely in a car engine. Assume petrol is all octane ( $C_8H_{18}$ ) and burns according to the equation:



- 7 Butane is used as the fuel in some portable camping stoves. It is a liquid when stored under pressure in a butane cylinder but vaporises when the valve is opened. Combustion of butane is represented by the equation:



- a How much energy is produced when 10.0 g of butane burns completely?  
b How much energy is produced when 0.100 L of butane, measured at SLC, burns completely?
- 8 Methane and methanol both burn in air according to the thermochemical equations:



- a If 1 mole of methane and 1 mole of methanol are completely burned in separate experiments, which experiment will release the most energy?  
b If each of the above reactions is used to produce 2000 kJ of energy, which fuel will release the most carbon dioxide?



## 3.2 Determination of limiting reactants or reagents

In the previous section, you learnt how to calculate the amount of product that could be formed, or energy that could be released, given a specified amount of a fuel. It was assumed for each combustion reaction that there was sufficient oxygen present for the fuel to be completely consumed.

In this section, you will learn an approach that you can take when it is not immediately obvious if a reactant or **reagent** is completely consumed in a reaction. A reactant is a starting material that undergoes change during a chemical reaction whereas a reagent is a substance added to a system to cause a chemical reaction, such as a catalyst. Both reactants and reagents can be a limiting substance in a reaction.

In these scenarios, you are given quantities of both reactants present. Before you can calculate the amount of product formed, you will need to work out which reactant is completely consumed.

### LIMITING AND EXCESS REACTANTS

When two reactants are mixed to create a chemical reaction, it is possible they will be combined in just the right mole ratio as indicated in the equation for each to be completely consumed. However, it is more likely that they are not present in exactly the right mole ratio, meaning one of the reactants will be used up before the other. Some of the other reactant will remain unreacted and be left over, or **in excess**, once the reaction has ceased.

To illustrate this situation simply, consider a problem in which you have been given some skateboard decks and wheels and you want to make as many complete skateboards as you can. As shown in Figure 3.2.1, a complete skateboard is made up of one deck and four wheels.



FIGURE 3.2.1 Construction of a complete skateboard requires one deck and four wheels.

If you were given two decks and ten wheels as shown in Figure 3.2.2, how many complete skateboards can you make from these materials?



FIGURE 3.2.2 When provided with ten wheels and two skateboard decks, how many complete skateboards can be made?

The answer is that you could make two complete skateboards and there would be two wheels left over (Figure 3.2.3).



FIGURE 3.2.3 When supplied with two decks and ten wheels the maximum number of skateboards that can be made is two. There will be two wheels that are not used.

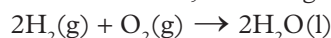
We can see that the number of skateboards that could be made was limited by the number of decks available. The decks were the limiting factor. The wheels were not completely used up and can be said to be in excess.

A similar situation arises in chemical reactions if the quantities of reactants present do not match the exact same ratio as that shown in the equation for the reaction.

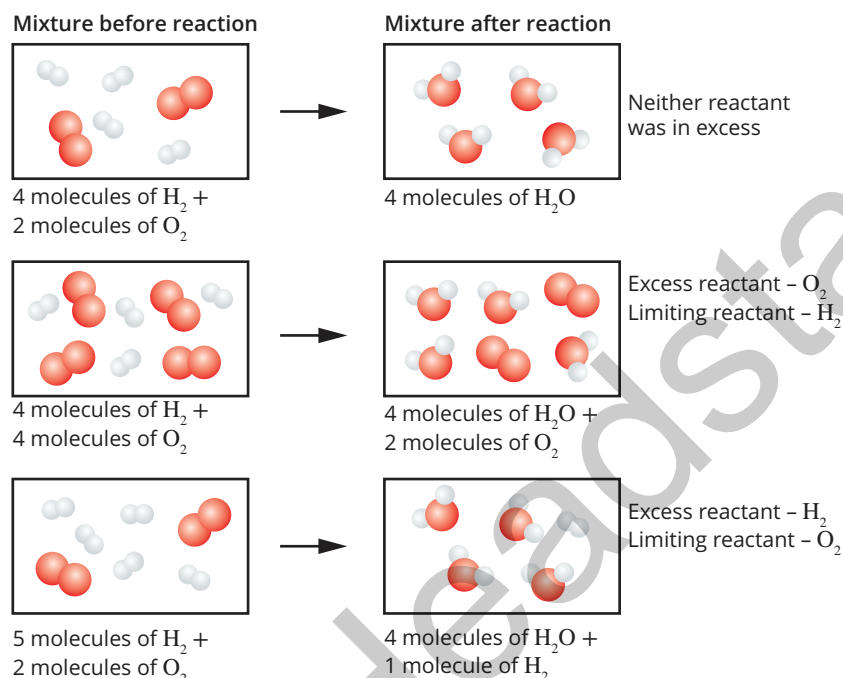
When this happens:

- the reactant that is completely consumed is called the **limiting reactant**
- the reactant that is not completely consumed is called the **excess reactant**.

Figure 3.2.4 shows three different scenarios for the reaction in which hydrogen gas and oxygen gas combine to form water, according to the equation:



Each of the diagrams provides examples to illustrate the concepts of limiting and excess reactants.



**FIGURE 3.2.4** Different scenarios showing the concept of limiting reactant for the reaction  $2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\text{l})$

Note that in each of the examples shown in Figure 3.2.4, the amount of product formed is determined by the amount of the limiting reactant present in the reaction mixture. The amount of product formed cannot be determined from the amount of excess reactant. This means it is essential to always identify the limiting reactant for use in stoichiometry.

### Steps in solving stoichiometry problems involving excess reactants

When attempting to solve a problem in which you are required to work out the limiting reactant and use stoichiometry to calculate the amount of product, there are three main steps:

- 1 Calculate the number of moles of each reactant.
- 2 Identify which reactant is the limiting reactant.
- 3 Use the amount of limiting reactant to determine the amount of product formed.

Each step is illustrated in Worked examples 3.2.1 and 3.2.2.

**i** In a chemical reaction, the limiting reactant is the reactant that is completely consumed in the reaction.

### CHEMFILE

#### Excess reactants in the chemical industry

Many chemical reactions are used in the chemical industry to make useful products from soap to pesticides and pharmaceuticals. The design of these processes will consider cost, product yield and minimising waste. If, in a particular chemical reaction, one reactant is more expensive than another, it will be important to ensure the expensive reactant is fully utilised. One way to assist with this is making sure the other reactant is in excess. Stoichiometric calculations help to determine how much reactant is required to ensure it is in excess. However, having an excess reactant may cause the disadvantage of requiring that reactant to be recovered and recycled.



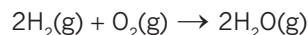
A chemist designing an industrial manufacturing process will consider which reactants or reagents to have in excess to maximise yield while minimising costs and waste.

**i** The amount of the limiting reactant must always be used to determine the amount of product that will be formed.

### Worked example 3.2.1

#### IDENTIFYING LIMITING AND EXCESS REACTANTS

A gaseous mixture of 25.0 g of hydrogen gas ( $\text{H}_2$ ) and 150 L of oxygen gas at SLC are mixed and ignited. The equation for the reaction is:



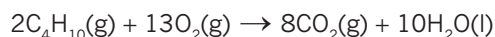
- a** Identify which reactant is the limiting reactant.  
**b** Calculate the amount, in mol, of the excess reactant that remains unreacted.

Thinking	Working
<p><b>a</b> Calculate the number of moles of each reactant using <math>n = \frac{m}{M}</math> or <math>n = \frac{V}{V_m}</math> or <math>n = \frac{pV}{RT}</math>. The formula chosen will depend on the quantity and conditions provided for each reactant.</p>	$n(\text{H}_2) = \frac{m}{M}$ $= \frac{25.0}{2.0}$ $= 12.5 \text{ mol}$ $n(\text{O}_2) = \frac{V}{V_m}$ $= \frac{150}{24.8}$ $= 6.05 \text{ mol}$
Choose one reactant and use the coefficients in the equation to find the amount of the other reactant needed for it to completely react.	$n(\text{H}_2) \text{ needed to react} = 2 \times n(\text{O}_2)$ $= 2 \times 6.05$ $= 12.1 \text{ mol}$
Compare the values for the amount required of the second reactant and the amount actually present to determine which is the limiting factor.	12.1 mol of $\text{H}_2$ is needed for all of the $\text{O}_2$ to react. There is 12.5 mol of $\text{H}_2$ present. The $\text{H}_2$ is in excess. $\text{O}_2$ is the limiting reactant.
<b>b</b> Use the mole ratio and amount of limiting reactant to determine the amount of excess reactant involved in the reaction.	$n(\text{O}_2) \text{ that reacted} = 6.05 \text{ mol}$ $\frac{n(\text{H}_2)}{n(\text{O}_2)} = \frac{2}{1}$ $n(\text{H}_2\text{O}) = 2 \times 6.05$ $= 12.1 \text{ mol}$
Determine the amount of excess reactant leftover after the limiting reactant has been consumed.	$n(\text{H}_2) \text{ in excess} = n(\text{H}_2) \text{ initially} - n(\text{H}_2) \text{ reacted}$ $= 12.5 - 12.1$ $= 0.4 \text{ mol}$

### Worked example: Try yourself 3.2.1

#### IDENTIFYING LIMITING AND EXCESS REACTANTS

8.00 g of butane ( $\text{C}_4\text{H}_{10}$ ) is burned completely in 20.0 L of oxygen at SLC. The equation for the reaction is:



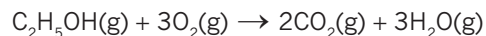
- a** Which reactant is the limiting reactant?  
**b** Calculate the amount, in mol, of the excess reactant that remains unreacted.



### Worked example 3.2.2

#### STOICHIOMETRY CALCULATIONS INVOLVING LIMITING AND EXCESS REACTANTS

100.0 g of ethanol ( $\text{C}_2\text{H}_5\text{OH}$ ) burns in 250.0 g of oxygen gas. The equation for the reaction is:



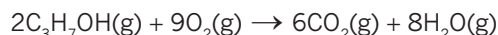
Calculate the volume, in L, of carbon dioxide formed at 100 kPa and 15°C.

Thinking	Working
Calculate the number of moles of each reactant using $n = \frac{m}{M}$ or $n = \frac{V}{V_m}$ or $n = \frac{pV}{RT}$ . The formula chosen will depend on the quantity and conditions provided for each reactant.	$n(\text{C}_2\text{H}_5\text{OH}) = \frac{m}{M}$ $= \frac{100}{46.0}$ $= 2.18 \text{ mol}$ $n(\text{O}_2) = \frac{m}{M}$ $= \frac{250.0}{32.0}$ $= 7.81 \text{ mol}$
Identify the limiting reactant.	$n(\text{O}_2) \text{ needed to react} = 3 \times n(\text{C}_2\text{H}_5\text{OH})$ $= 3 \times 2.18$ $= 6.54 \text{ mol}$ <p>As there is 7.81 mol of <math>\text{O}_2</math> present, the <math>\text{O}_2</math> is in excess. Ethanol is the limiting reactant.</p>
Find the mole ratio: $\frac{\text{coefficient of unknown}}{\text{coefficient of known (limiting reactant)}}$	$\frac{n(\text{CO}_2)}{n(\text{C}_2\text{H}_5\text{OH})} = \frac{2}{1}$
Calculate the number of moles of the unknown substance using: $n(\text{unknown}) = \text{mole ratio} \times n(\text{known})$	$n(\text{CO}_2) = \frac{2}{1} \times 2.18$ $= 4.36 \text{ mol}$
Calculate the required quantity of the unknown substance using $m = n \times M$ , $V = n \times V_m$ or $V = \frac{nRT}{p}$ as appropriate. The choice will depend on the required quantity and provided conditions.	<p>The volume of a gas at non-standard conditions needs to be calculated, so the formula chosen is <math>V = \frac{nRT}{p}</math></p> $V(\text{CO}_2) = \frac{4.36 \times 8.31 \times 288}{100}$ $= 104 \text{ L}$

### Worked example: Try yourself 3.2.2

#### STOICHIOMETRY CALCULATIONS INVOLVING LIMITING AND EXCESS REACTANTS

150 g of propanol ( $\text{C}_3\text{H}_7\text{OH}$ ) burns in 200.0 g of oxygen gas. The equation for the reaction is:



Calculate the volume, in L, of carbon dioxide formed at 120 kPa and 20°C.

## 3.2 Review



### SUMMARY

- If quantities of more than one reactant or reagent are provided for a chemical reaction, the limiting reactant or reagent must be identified and used in stoichiometry calculations.
- A reactant is a starting material that undergoes change during a chemical reaction, whereas a reagent is a substance added to a system to cause a chemical reaction, such as a catalyst.
- The limiting reactant or reagent is the reactant or reagent that is completely consumed in the reaction.
- The reactant or reagent that is not completely used up and has some remaining when the reaction has stopped is said to be in excess.

### KEY QUESTIONS

#### Knowledge and understanding

- 1 Iron reacts with oxygen gas to form iron oxide. The balanced equation for the reaction is:
- $$4\text{Fe(s)} + 3\text{O}_2\text{(g)} \rightarrow 2\text{Fe}_2\text{O}_3\text{(s)}$$
- In a particular reaction, 25 g of iron reacts with 60 g of oxygen gas. List the following steps in order to determine the mass of iron oxide that will form.
- Calculate the mass of iron oxide that forms.
  - Refer to the balanced equation.
  - Calculate the number of moles of iron oxide that forms.
  - Calculate the number of moles of iron and oxygen.
  - Use mole ratios to determine which reactant is limiting.
- 2 In three different experiments, different amounts of propane ( $\text{C}_3\text{H}_8$ ) and oxygen undergo combustion according to the equation:
- $$\text{C}_3\text{H}_8\text{(g)} + 5\text{O}_2\text{(g)} \rightarrow 3\text{CO}_2\text{(g)} + 4\text{H}_2\text{O(g)}$$
- This table shows the amount of reactants and products in each experiment. Complete the table to indicate the amount of each product remaining at the end of the reaction.

Propane molecules available	Oxygen molecules available	Carbon dioxide molecules produced	Propane molecules in excess	Oxygen molecules in excess
2	15			
300	1200			
2.5 mol	10 mol			

#### Analysis

- 3 Fluorine gas reacts with solid phosphorus to form phosphorus trifluoride ( $\text{PF}_3$ ) according to the following reaction:
- $$3\text{F}_2\text{(g)} + 2\text{P(s)} \rightarrow 2\text{PF}_3\text{(g)}$$
- 42.0 g of fluorine gas is reacted with 29.0 g of phosphorus. Calculate the mass of phosphorus trifluoride that will form in the reaction.
- 4 7.0 g of ethanol ( $\text{C}_2\text{H}_5\text{OH}$ ) and 30.0 g of  $\text{O}_2$  are mixed in a closed vessel and allowed to react according to this equation:
- $$\text{C}_2\text{H}_5\text{OH(l)} + 3\text{O}_2\text{(g)} \rightarrow 2\text{CO}_2\text{(g)} + 3\text{H}_2\text{O(g)}$$
- Which reactant is in excess and by what mass, in g?
  - What mass of  $\text{CO}_2$  forms?
  - What mass of  $\text{H}_2\text{O}$  forms?
  - What is the total mass of the mixture after the reaction is complete?
- 5 23.0 g of ethanol ( $\text{C}_2\text{H}_5\text{OH}$ ) is reacted with 10.0 L of oxygen at SLC. Calculate the volume of carbon dioxide produced in this reaction.
- The equation for the reaction is:
- $$\text{C}_2\text{H}_5\text{OH(g)} + 3\text{O}_2\text{(g)} \rightarrow 2\text{CO}_2\text{(g)} + 3\text{H}_2\text{O(l)}$$
- 6 8.50 g of propanol ( $\text{C}_3\text{H}_7\text{OH}$ ) is reacted with 20.0 L of oxygen at 120 kPa and  $30^\circ\text{C}$ . Calculate the volume of carbon dioxide produced in this reaction.
- The equation for the reaction is:
- $$2\text{C}_3\text{H}_7\text{OH(g)} + 9\text{O}_2\text{(g)} \rightarrow 6\text{CO}_2\text{(g)} + 8\text{H}_2\text{O(l)}$$

## 3.3 Calculating heat energy released

Knowing the energy released by similar quantities of different fuels helps you to compare fuels and determine their suitability for specific purposes. For example, the fuel used to power an aeroplane (Figure 3.3.1) is different from the fuel used to power a car or bus.

In Chapter 2 you learnt how thermochemical equations include a  $\Delta H$  value, which shows the amount of heat released when a fuel undergoes combustion. The value of  $\Delta H$  is based on the stoichiometric ratios in the equation.

In this section, you will learn how knowledge of the specific heat capacity of water can be used to obtain an experimental estimate of **heat of combustion**, as well as how the heat of combustion can be used to write thermochemical equations.

### SPECIFIC HEAT CAPACITY OF WATER

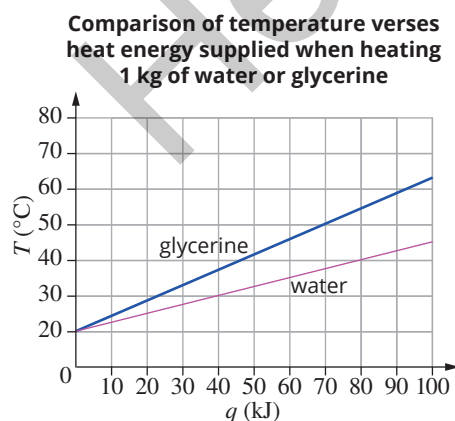
The **specific heat capacity** of a substance is a measure of the amount of energy (usually in joules) needed to increase the temperature of a specific quantity of that substance (usually 1 gram) by  $1^{\circ}\text{C}$ .

Specific heat capacity is given the symbol  $c$  and is usually expressed in joules per gram per degrees Celsius, i.e.  $\text{J g}^{-1}\text{ }^{\circ}\text{C}^{-1}$ . It can also be expressed in joules per grams per kelvin, i.e.  $\text{J g}^{-1}\text{ K}^{-1}$  (an increase of  $1^{\circ}\text{C}$  is the same as an increase of 1 K).

The specific heat capacities of some common substances are listed in Table 3.3.1. You can see that the value for water is relatively high.

The specific heat capacity of a substance reflects the types of bonds holding the molecules, ions or atoms together in the substance. Water has a specific heat capacity of  $4.18\text{ J g}^{-1}\text{ }^{\circ}\text{C}^{-1}$ . This means that 4.18 joules of heat energy are needed to increase the temperature of 1 gram of water by  $1^{\circ}\text{C}$ . This relatively high value is due to the hydrogen bonds between the water molecules. The higher the specific heat capacity, the more effectively a material stores heat energy.

When a substance is being heated, its temperature rises. The temperature of 1 g of water increases by  $1^{\circ}\text{C}$  when it is supplied with 4.18 J of heat energy. In comparison, 2.43 J of heat energy is required to increase the temperature of 1 g of glycerine by  $1^{\circ}\text{C}$ . The effect of the different specific heat capacities of water and glycerine on their temperatures when heated can be seen in Figure 3.3.2.



**FIGURE 3.3.2** A comparison of the effect of the different specific heat capacities of water and glycerine on the increase in temperature ( $T$ ). Water has a very high heat capacity, so it requires more heat energy ( $q$ ) to increase its temperature by  $1^{\circ}\text{C}$ .



**FIGURE 3.3.1** F15 jets being refuelled by a Boeing 707. The type of fuel suitable for use in aircraft is different from the type of fuel suitable for use in cars or buses.

**TABLE 3.3.1** Specific heat capacities of some common liquids

Substance	Specific heat capacity ( $\text{J g}^{-1}\text{ }^{\circ}\text{C}^{-1}$ )
water	4.18
glycerine	2.43
ethanol	2.46
hexane	2.26
olive oil	1.97
paraffin oil	2.13



## TRANSFERRING HEAT ENERGY TO WATER

The specific heat capacity of water can be used to calculate the heat energy in joules needed to increase the temperature of a given mass of water by a particular amount. Heat energy is given the symbol  $q$ , as seen in Figure 3.3.2 on the previous page.

The heat energy transferred to the mass of water can be calculated by using measurements of the:

- initial temperature of the water
- highest temperature of the water
- mass of water.

The relationship used to calculate the energy that has been transferred to the water is:

Heat energy = mass of water  $\times$  specific heat capacity  $\times$  temperature change

Using symbols, the equation can be written as:

$$q = m \times c \times \Delta T \quad \text{or} \quad q = mc\Delta T$$

where  $q$  is the amount of heat energy (in J),  $m$  is the mass (in g),  $c$  is the specific heat capacity of the water ( $4.18 \text{ J g}^{-1} \text{ }^\circ\text{C}^{-1}$ ) and  $\Delta T$  is the temperature change (in  $^\circ\text{C}$  or K):

$$\Delta T = \Delta T_{\text{final}} - \Delta T_{\text{initial}}$$

Recall that the density of water is  $1.0 \text{ g mL}^{-1}$ , so 1 mL of water has a mass of 1 g. Sometimes the density of water is quoted as  $0.997 \text{ g mL}^{-1}$ , however, this value, to 3 significant figures, is only accurate at  $25^\circ\text{C}$ .

### Worked example 3.3.1

#### CALCULATING THE ENERGY REQUIRED TO HEAT A MASS OF WATER USING SPECIFIC HEAT CAPACITY

Calculate the heat energy, in kJ, needed to increase the temperature of 500 mL of water by  $15.0^\circ\text{C}$ .

Thinking	Working
Change the volume of water, in mL, to mass of water, in g. Remember that the density of water is $1.0 \text{ g mL}^{-1}$ so 1 mL of water has a mass of 1 g.	500 mL of water has a mass of 500 g.
Find the specific heat capacity of water from the data in Table 3.3.1.	The specific heat capacity of water is $4.18 \text{ J g}^{-1} \text{ }^\circ\text{C}^{-1}$ .
To calculate the quantity of heat energy in joules, use the formula: $q = mc\Delta T$	$q = 500 \times 4.18 \times 15.0$ $= 3.14 \times 10^4 \text{ J}$
Express the quantity of energy in kJ. Remember that to convert from J to kJ, you multiply by $10^{-3}$ .	$q = 3.14 \times 10^4 \times 10^{-3}$ $= 31.4 \text{ kJ}$

### Worked example: Try yourself 3.3.1

#### CALCULATING THE ENERGY REQUIRED TO HEAT A MASS OF WATER USING SPECIFIC HEAT CAPACITY

Calculate the heat energy, in kJ, needed to increase the temperature of 375 mL of water by  $45.0^\circ\text{C}$ .

### Experimental determination of heat of combustion

When an **exothermic** chemical reaction, such as the combustion of a pure substance, is carried out underneath a container of water, such as a metal can or a test tube, some of the heat released by the combustion reaction is transferred to the water. If you measure the temperature change of the water, it can be used to determine the approximate amount of energy released by the substance.

**i** The heat energy required to increase the temperature of a given mass of water by a particular amount can be calculated using the equation:

$$q = mc\Delta T$$

#### CHEMFILE

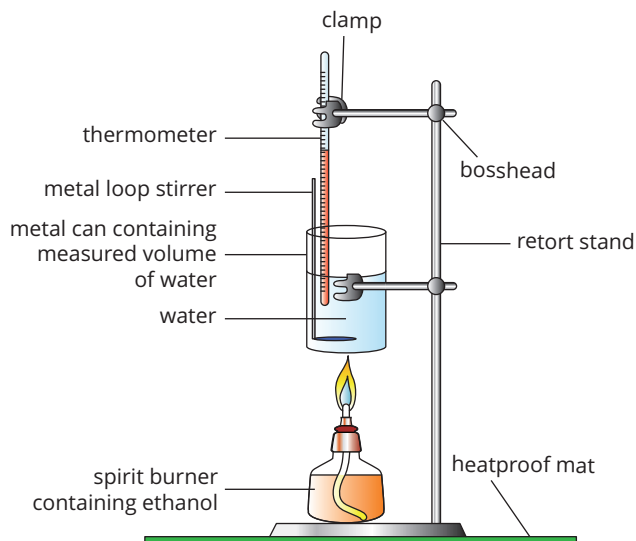
#### Hot beach, cool bay

The different heat capacities of water and sand are evident at the beach on a hot summer's day. As a cloudless day dawns, both sand and water sit under a hot sun, as seen in the figure below. Sand which is not in contact with water from waves or tides can heat very quickly to the point where it could cause burns to the bottom of your feet. In contrast, the water, even in rock pools where constant mixing is not occurring, remains cool and refreshing. The specific heat capacity of water is  $4.18 \text{ J g}^{-1} \text{ }^\circ\text{C}^{-1}$ , while that of sand is  $0.48 \text{ J g}^{-1} \text{ }^\circ\text{C}^{-1}$ , so water can absorb nearly 10 times as much energy as sand for the same temperature increase. Each gram of water will absorb 4.18 J before it increases by  $1^\circ\text{C}$ . Each gram of sand will absorb only 0.48 J before it increases by  $1^\circ\text{C}$ .



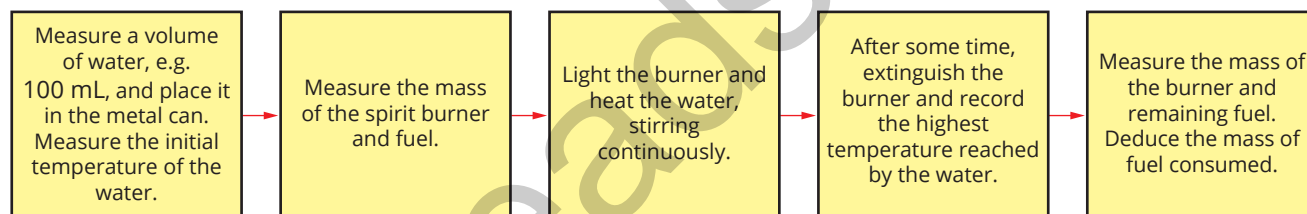
At this Mornington Peninsula beach, the sand on a sunny day can be extremely hot while the water remains cool.

An experimental arrangement for estimating the heat of combustion of an organic liquid, such as ethanol, is shown in Figure 3.3.3. This experimental determination is suitable for many organic liquids, such as alcohols and some alkanes and alkenes. Because these liquids release a lot of energy when they are burnt, some of them could be regarded as fuels.



**FIGURE 3.3.3** Apparatus for measuring heat of combustion of an organic liquid fuel (e.g. ethanol). A metal can containing a measured volume of water is held above the wick of a spirit burner.

Figure 3.3.4 summarises the steps followed in this experiment.



**FIGURE 3.3.4** Flow chart of the steps followed when using the specific heat capacity of water to determine the heat of combustion of a pure organic liquid fuel

Three key pieces of information collected from this procedure are the:

- mass of water
- change in temperature of the water,  $\Delta T$
- mass of organic liquid fuel consumed, from which the amount, in mol, of the substance can be calculated.

In Chapter 2 you learnt that the heat of combustion of a fuel can be determined when the energy released by that fuel is divided by the amount, in mol, of the fuel that is burnt. You also learnt that the heat of combustion of a fuel is the heat energy released when a specified amount of the substance burns completely and is a positive value, while the enthalpy of combustion of a fuel reflects its exothermic nature with a negative sign and has the symbol,  $\Delta H_c$ .

The heat of combustion can be calculated using the equation:

$$\text{heat of combustion} = \frac{q}{n}$$

where  $q$ , the energy absorbed by the water in the can, is calculated using the equation:

$$q = mc\Delta T$$

and  $n$  is the amount of fuel, in this case the pure organic liquid, that has been burnt.

**i** Recall that the heat of combustion of a fuel is reported as a positive value due to its definition, while the enthalpy of combustion,  $\Delta H_c$  is written in a thermochemical equation and is negative.

This calculation of heat of combustion assumes that all the energy is transferred from the burning pure organic liquid fuel to the mass of water. Substantial heat losses occur when the experiment described above is performed, so the values calculated for the heat of combustion will be less than the actual values. At best, this experiment could only be used to obtain an estimate of the heat of combustion or perhaps to compare the energy released by two or more pure substances, fuels or foods.

This is shown in Worked example 3.3.2.

### Worked example 3.3.2

#### CALCULATING THE HEAT OF COMBUSTION OF A PURE ORGANIC LIQUID FUEL FROM EXPERIMENTAL DATA

0.355 g of methanol ( $\text{CH}_3\text{OH}$ ) undergoes complete combustion in a spirit burner. The heat energy released is used to heat 100 g of water. The temperature of the water rose from  $20.24^\circ\text{C}$  to  $37.65^\circ\text{C}$ . Calculate the heat of combustion of methanol in  $\text{kJ mol}^{-1}$  and write the thermochemical equation for the reaction.

Thinking	Working
Calculate the temperature change of the water.	$\Delta T = 37.65 - 20.24$ $= 17.41^\circ\text{C}$
Use the specific heat capacity of water to determine the energy used to heat the water. Use the formula: $q = mc\Delta T$	$q = 100 \times 4.18 \times 17.41$ $= 7277 \text{ J}$
Express the quantity of energy in kJ. Remember that to convert from J to kJ, you multiply by $10^{-3}$ .	$q = 7277 \times 10^{-3}$ $= 7.277 \text{ kJ}$
Calculate the amount, in moles, of methanol using the formula: $n = \frac{m}{M}$	$n = \frac{0.355}{32.0}$ $= 0.0111 \text{ mol}$
Determine the heat of combustion of methanol, in $\text{kJ mol}^{-1}$ . Heat of combustion = $\frac{\text{heat energy released by sample}}{\text{amount of sample (in mol)}}$ Or heat of combustion = $\frac{q}{n}$	Heat of combustion $= \frac{7.277}{0.0111}$ $= 656 \text{ kJ mol}^{-1}$
The thermochemical equation for this reaction is a balanced equation, including the enthalpy of combustion, $\Delta H$ , which has a negative sign.	$\text{CH}_3\text{OH(l)} + \frac{3}{2}\text{O}_2\text{(g)} \rightarrow$ $\text{CO}_2\text{(g)} + 2\text{H}_2\text{O}; \Delta H = -656 \text{ kJ}$ or $2\text{CH}_3\text{OH(l)} + 3\text{O}_2\text{(g)} \rightarrow$ $2\text{CO}_2\text{(g)} + 4\text{H}_2\text{O}; \Delta H = -1312 \text{ kJ}$

### Worked example: Try yourself 3.3.2

#### CALCULATING THE HEAT OF COMBUSTION OF A PURE ORGANIC LIQUID FUEL FROM EXPERIMENTAL DATA

0.295 g of ethanol ( $\text{C}_2\text{H}_5\text{OH}$ ) undergoes complete combustion in a spirit burner. The heat energy released is used to heat 100 g of water. The temperature of the water rose from  $19.56^\circ\text{C}$  to  $38.85^\circ\text{C}$ . Calculate the heat of combustion of ethanol in  $\text{kJ mol}^{-1}$  and write the thermochemical equation for the reaction.

**i** To calculate the heat of combustion of a fuel, the following two equations are used.  
 $q = mc\Delta T$   
heat of combustion =  $\frac{q}{n}$

## ESTIMATING THE ENERGY CONTENT OF A FUEL OR FOOD

Because foods and some fuels, such as diesel, are mixtures, rather than a pure substance, the heat energy released during the combustion of a known mass of these fuels or food is best described as an **energy content**, in  $\text{kJ g}^{-1}$ .

If the mass of fuel or food burned to produce this energy is measured, the energy content of the fuel or food can be calculated as follows:

$$\text{energy content} = \frac{\text{energy transferred to the water}}{\text{change in mass of the fuel/food during combustion}} = \frac{q}{\Delta m}$$

The mass of fuel or food burned is determined by measuring the initial mass of fuel or food, and then subtracting the final mass. It is not necessary to burn all the fuel or food sample, as long as the mass that is burned is determined.

In this type of experiment, heat loss is a consistent problem, so the values calculated for the energy content are an estimate only and will be less than the actual values.

Worked example 3.3.3 describes how to calculate energy content from simple experimental data.

### Worked example 3.3.3

#### ESTIMATING THE ENERGY CONTENT OF A SAMPLE OF FOOD

A 1.670 g sample of cheese biscuit was burned under a steel can containing 200 g of water. After the flame went out, the mass of the cheese biscuit was 0.300 g and the temperature of the water had risen by  $29.8^\circ\text{C}$ . Calculate the energy content of the biscuit in  $\text{kJ g}^{-1}$ .

Thinking	Working
Calculate the heat energy absorbed by the water in joules, using the formula: $q = mc\Delta T$	$q = mc\Delta T$ $= 200 \times 4.18 \times 29.8$ $= 24\,913 \text{ J}$
Express the quantity of energy in kJ. Remember that to convert from joules to kilojoules, you divide by $10^3$ or multiply by $10^{-3}$ .	$q = \frac{24\,913}{1000}$ $= 24.913 \text{ kJ}$
Calculate the mass of the food that was burned by subtracting the final mass from the initial mass: $\Delta m = \Delta m_{\text{initial}} - \Delta m_{\text{final}}$	$\Delta m = \Delta m_{\text{initial}} - \Delta m_{\text{final}}$ $= 1.670 - 0.300$ $= 1.370 \text{ g}$
Calculate the energy content of the food by dividing the energy transferred to the water by the change in mass during combustion: Energy content = $\frac{q}{\Delta m}$	Energy content = $\frac{q}{\Delta m}$ $= \frac{24.913}{1.370}$ $= 18.2 \text{ kJ g}^{-1}$ (3 significant figures)

### Worked example: Try yourself 3.3.3

#### ESTIMATING THE ENERGY CONTENT OF A SAMPLE OF FOOD

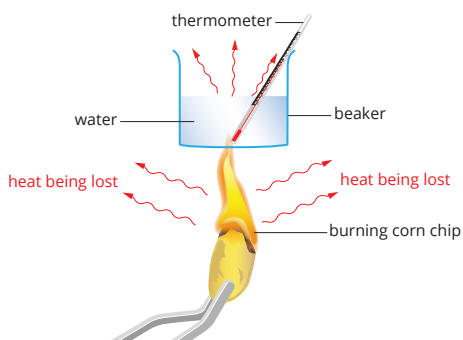
A 2.500 g sample of a corn chip was burned under a steel can containing 200 g of water. After the flame went out, the mass of the corn chip sample was 1.160 g and the temperature of the water had risen by  $35.0^\circ\text{C}$ . Calculate the energy content of the corn chip in  $\text{kJ g}^{-1}$ .

**i** To calculate the energy content of a food or fuel, use the equation

$$\text{energy content} = \frac{q}{m}$$

where  $q$  is the energy transferred to the water and  $\Delta m$  is the change in the mass of the food or fuel.





**FIGURE 3.3.5** Heat is lost to the surroundings when a sample of food is burned to heat some water.

## Heat loss

When energy is transferred from burning a sample of fuel or food, across an open space, heat is lost to the surroundings such as the air around the burning fuel. Similarly, if there is no lid on a container of water, heat will be lost from the surface of the water. This heat loss is illustrated in Figure 3.3.5.

When some heat energy from the burning fuel or food is transferred to the surrounding air, the temperature of the water does not increase as much as it would if all the energy was used to heat the water. A lower change in temperature,  $\Delta T$ , of the water results in a lower energy value,  $q$ .

There are several ways to reduce heat loss during the experiment shown in Figure 3.3.5, including:

- putting a lid on the container holding the water
- insulating the beaker of water (with flameproof material)
- placing insulation around the burning fuel, although sufficient oxygen must reach the fuel for combustion to be complete.

As mentioned earlier, this loss of heat energy represents a significant systematic error in the calculation of heat of combustion by this experimental method. More accurate determinations of energy content require a more sophisticated piece of equipment called a bomb calorimeter.

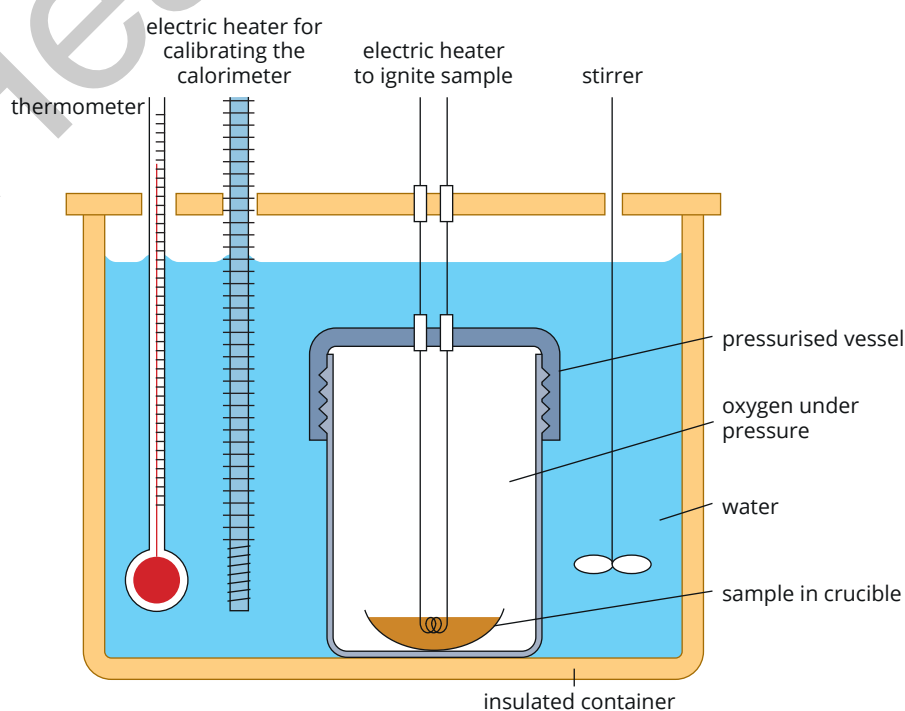
## Bomb calorimetry

Figure 3.3.6 shows the components of a **bomb calorimeter**, which is a piece of equipment used for measuring the energy released by combustion reactions that involve gaseous reactants or products. The reaction vessel in a bomb calorimeter is designed to withstand the high pressures that may build up during reactions. An actual bomb calorimeter can be seen in Figure 3.3.7.

The energy content of fuels is measured by burning them in a bomb calorimeter. Sufficient oxygen is required to completely combust the fuel so that all the available energy is released. Insulation around the calorimeter prevents heat escaping and the change in temperature is measured with a thermometer. The stirrer ensures that the temperature of the water is uniform. You will learn more about calorimetry in Section 3.4.



**FIGURE 3.3.7** Bomb calorimeter temperature changes being observed by a laboratory technician



**FIGURE 3.3.6** A bomb calorimeter used for reactions that involve gases, including the combustion of fuels

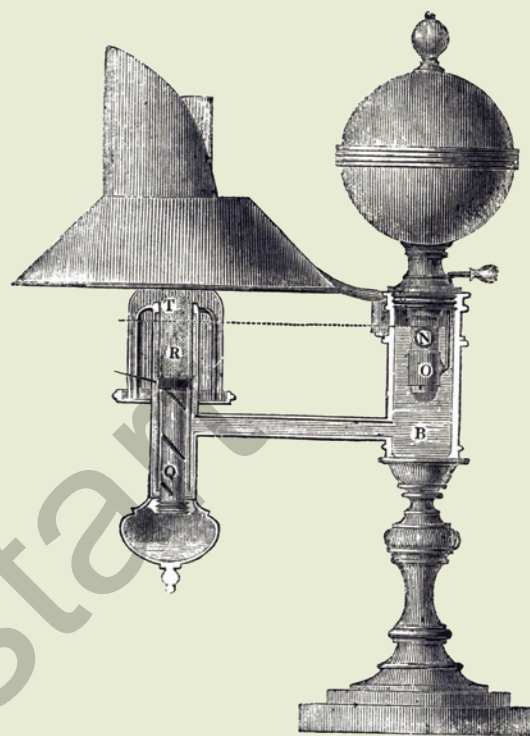
## The energy of candlelight

Little is known about the origins of candles, but it is thought that they have been around for over 5000 years of human civilisation. Candle wax is made of large molecules composed of carbon and hydrogen atoms, such as eicosane,  $C_{20}H_{42}$ , and it has an approximate energy content of  $47.1 \text{ kJ g}^{-1}$ . Candles were used by the ancient Romans as early as 500 BCE. However, Romans used oil lamps in which olive oil was burnt as their main light source. The burning of whale oil was common practice in the use of lamps from the sixteenth to the nineteenth centuries. Whale oil is a mixture containing large quantities of oleic acid which has an energy content of  $39 \text{ kJ g}^{-1}$ .

One such lamp was developed in 1780 (see Figure 3.3.8) by François-Pierre-Amédée (also known as Ami) Argand, a student of Antoine Lavoisier. These lamps were considered to have about the same light output as 6 to 8 candles. In fact, a 'candela' or candle power is a measure of the intensity of light, now known as lumens. A modern 10 Watt LED globe has an output of over 300 lumens. Argand further developed his lamp into a burner that was used until the development of the Bunsen burner.

Cheaper alternatives to whale oil were available at that time, but they did not burn as cleanly, producing large quantities of carbon soot. Demand for whale oil vastly increased in the seventeenth century and whale hunting reached its peak in the 1820s. After that time there was a rapid decline in use of whale oil, as cheaper and cleaner-smelling combustible materials became more common. Camphine, also known as 'oil of turpentine', a mixture of turpentine and alcohol, was used until kerosene rapidly dominated the market from the 1860s, making the use of camphine insignificant by 1866.

Now, in the twenty-first century, whale hunting is banned across most of the globe, and we use electricity to power our lights. Just as humans have moved away from the use of whale oil for lighting, we have the potential to explore more renewable fuel sources for our energy needs, so that we can preserve our environment.



**FIGURE 3.3.8** An engraving depicting a sectional view of an Argand lamp, a kind of oil lamp, that was invented and patented in 1780 by Ami Argand

### Analysis

- 1 Write a balanced chemical reaction for the complete combustion of eicosane,  $C_{20}H_{42}$ .
- 2 What is occurring when a fuel, such as oil, is burning? Refer to states of matter, energy absorbed and released, and bonding in your answer.
- 3 A student decided to use a candle made only of eicosane to heat a beaker containing 100 mL of water. Assuming that there is no heat loss, calculate the mass of candle that would be used to heat the water from  $18.0^\circ\text{C}$  to  $100.0^\circ\text{C}$ .

## 3.3 Review



### SUMMARY

- The specific heat capacity of a substance measures the quantity of energy (in joules) needed to increase the temperature of a specified quantity of that substance (usually 1 gram) by 1°C.
- The specific heat capacity of water is 4.18 J g<sup>-1</sup> °C<sup>-1</sup>.
- The heat energy required to increase the temperature of a given mass of water by a particular amount can be calculated using the equation:

$$q = mc\Delta T$$

where  $q$  is heat energy (in J),  $m$  is mass of water to be heated (in g),  $c$  is the specific heat capacity (in J g<sup>-1</sup> °C<sup>-1</sup>) and  $\Delta T$  is the temperature change (in °C).

- In the experimental determination of the heat energy released in the combustion of a fuel, the specific heat capacity of water is used to determine the heat energy absorbed by a measured mass of water placed above the burning fuel.
- The heat of combustion of a pure substance (in kJ mol<sup>-1</sup>) can be calculated using the experimental results of the energy absorbed by the measured mass of water (using  $q = mc\Delta T$ ) and the amount of the substance (in mol) whose combustion released that energy using the equation:

$$\text{heat of combustion} = \frac{q}{n}$$

where  $q$  is heat energy (in kJ) and  $n$  is the amount (in mol) of fuel that was burnt.

- Because foods and some fuels are mixtures, their energy content is measured in kJ g<sup>-1</sup>. This is determined using the equation energy content =  $\frac{q}{m}$ , where energy is measured in kJ.
- Calculations of heat of combustion of a pure substance or energy content of a food or fuel from experimental data can be very inaccurate due to extensive heat loss to the surroundings during the experiment.
- A bomb calorimeter is an insulated container in which a sealed, oxygen-filled reaction vessel is surrounded by a known volume of water.
- Combustion reactions are carried out in the reaction vessel and the heat from the reaction is transferred to the surrounding water.

### KEY QUESTIONS

#### Knowledge and understanding

- 1 Calculate the energy, in kJ, required to increase the temperature of:
  - a 100 g of water by 25.1°C
  - b 500 g of water from 17.0°C to 80.0°C
  - c 1.50 kg of water from 20.0°C to 30.0°C.
- 2 A beaker containing 200 g of water at a temperature of 21.0°C is heated with 10.0 kJ of energy. Calculate the temperature reached by the water.

#### Analysis

- 3 A temperature rise of 1.78°C was observed when  $1.00 \times 10^{-3}$  mol of propane gas (C<sub>3</sub>H<sub>8</sub>) was burnt and used to heat 300 g of water. Calculate the heat of combustion of propane in kJ mol<sup>-1</sup>, using this experimental data.

- 4 A temperature rise of 11.5°C was observed when 0.500 g of butane gas (C<sub>4</sub>H<sub>10</sub>) was burnt and used to heat 500 g of water. Calculate the heat of combustion, in kJ g<sup>-1</sup>, for butane, assuming all the heat released is used to heat the water.
- 5 The heat of combustion of methane (CH<sub>4</sub>) is 890 kJ mol<sup>-1</sup>. Determine the mass of methane, in g, which needs to be burnt to heat 500 g of water from 20.0°C to boiling. Assume all the heat released is used to heat the water.
- 6 If the heat of combustion of a sample of sub-bituminous coal is 20.7 kJ g<sup>-1</sup>, what is the mass of coal that is burned when 62.7 kJ of energy is transferred to a container of water, assuming all the energy from the burning coal is absorbed by the water?

## 3.4 Solution calorimetry

**Calorimetry** is the experimental method of measuring the heat energy released or absorbed by a chemical reaction or physical process, such as by the combustion of a fuel.

Energy changes that occur during chemical and physical changes are measured with a device called a calorimeter. **Calorimeters** are constructed in such a way that the energy losses that occur in the simple experimental apparatus described earlier are minimised, enabling more accurate results. In a calorimeter, almost all of the heat energy released or absorbed is transferred directly to or from a measured volume of water.

Two types of calorimeters are designed for measuring the energy changes in different types of reactions.

- In solution calorimeters, the reaction takes place in a solution.
- In bomb calorimeters, the reaction takes place in a sealed bomb vessel. This was described briefly in Section 3.3.

### USING A SOLUTION CALORIMETER

Energy changes for reactions that occur in solution can be measured with a **solution calorimeter**. A solution calorimeter may be as simple as a polystyrene foam coffee cup with a lid, as shown in Figure 3.4.1.

The insulation provided by the polystyrene foam prevents the transfer of heat to or from the surroundings of the calorimeter. The reaction is carried out in the calorimeter with an accurately known volume of water. The initial and final temperatures are measured and recorded, as are the amounts of reactants used.

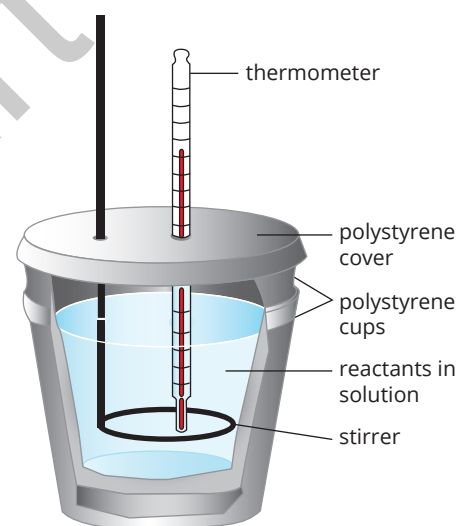
If the temperature of the water in the calorimeter increases, the reaction occurring in the calorimeter is an exothermic reaction. The reaction has released heat energy, the water in the calorimeter has absorbed that energy and the temperature of the water has increased.

Similarly, if the temperature of the water in the calorimeter decreases, the reaction occurring in the calorimeter has absorbed energy from the water. In this case, the reaction is an **endothermic** reaction.

A coffee cup solution calorimeter has some limitations. The polystyrene container absorbs some heat, so the temperature change is lower than it would otherwise be and the calculated value for the heat released or absorbed by the reaction is lower than it should be.

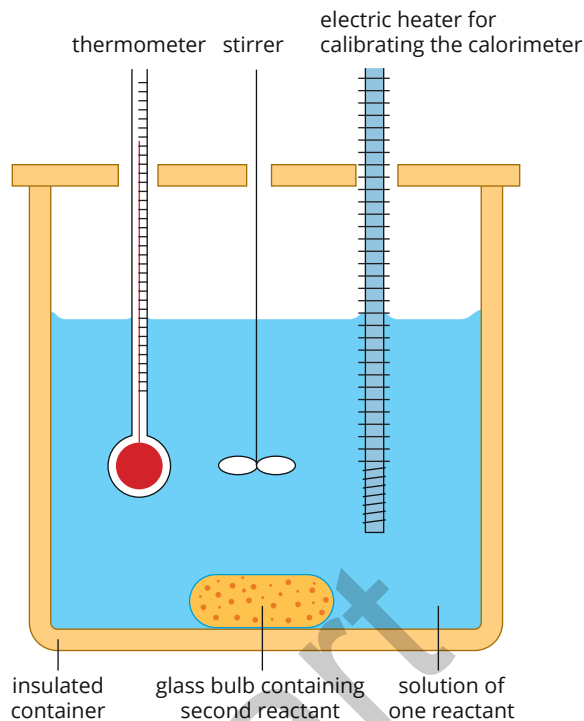
A solution calorimeter cannot be used to measure the energy content of fuels or foods because the reaction in this case is a combustion reaction in which the fuel burns in oxygen. However, a solution calorimeter can be used to determine the energy change that occurs when a substance such as glucose dissolves in water. Solution calorimetry is used in the laboratory to determine the enthalpy changes that occur when acids react with bases, metals react with acids, and solids dissolve in water.

The construction of a laboratory solution calorimeter is shown in Figure 3.4.2 on the following page. The stirrer is used to ensure the temperature of the water is uniform. If the temperature of the solution is not uniform, then the change in temperature measured as a result of the dissolution will not be accurate. The use of the electrical heater for calibrating the calorimeter is described in the next section.



**FIGURE 3.4.1** A simple 'coffee-cup' solution calorimeter





**FIGURE 3.4.2** A solution calorimeter: breaking the glass bulb starts the reaction.

## CALIBRATION OF CALORIMETERS

When a reaction takes place in a calorimeter, the heat change causes a rise or fall in the temperature of the contents of the calorimeter. Although the formula  $q = mc\Delta T$  can be used to determine the relationship between heat energy and temperature change, the heat losses that occur can lead to inaccurate results.

For more accurate measurements of the heat produced in a reaction, you have to first determine how much energy is required to change the temperature of the water by  $1^\circ\text{C}$  for the particular calorimeter in use. This is known as the **calibration factor** (CF) of the calorimeter. For energy measured in joules, the calibration factor has the unit  $\text{J } ^\circ\text{C}^{-1}$ . Once the calibration factor of a calorimeter is known, it is said to be **calibrated**.

Both solution and bomb calorimeters are usually calibrated before use.

### Electrical calibration of calorimeters

A calorimeter can be calibrated using the electric heater to release a known quantity of thermal energy and measuring the resultant rise in temperature of the water in the calorimeter.

The thermal energy released when an electric current passes through the heater can be calculated from the formula:

$$\text{energy (in joules)} = \text{voltage (volts)} \times \text{current (amps)} \times \text{time (seconds)},$$

or

$$E = V \times I \times t$$

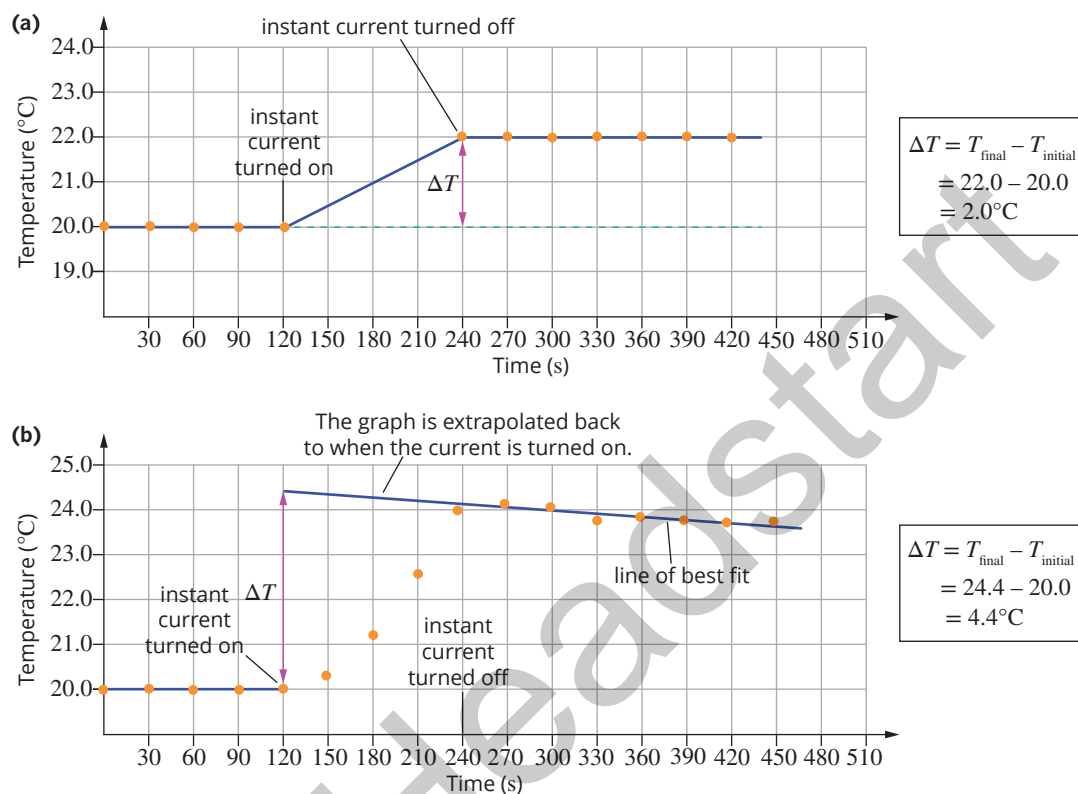
If the temperature change,  $\Delta T$ , caused by the addition of the known amount of energy is measured, the calibration factor for the calorimeter can be calculated from the expression:

$$CF = \frac{E}{\Delta T} = \frac{VIt}{\Delta T}$$

### Temperature–time graphs in calorimetry

Measuring the temperature change,  $\Delta T$ , that occurs during a calorimetry experiment is not always as simple as calculating the difference between the final temperature and the initial temperature of the water in the calorimeter. If a calorimeter is not perfectly insulated, it may slowly lose heat during and after the heater is operating. A more accurate determination of the temperature change ( $\Delta T$ ) can be achieved by plotting a graph of temperature against time before, during and after the calibration.

Two temperature–time graphs for the calibration of different calorimeters are shown in Figure 3.4.3. Figure 3.4.3a shows the results for a calorimeter with ‘perfect’ insulation and no heat loss. Figure 3.4.3b is typical of the results obtained using school laboratory calorimeters.



**FIGURE 3.4.3** Graphs of temperature in a solution calorimeter before, during and after calibration. (a) These results would be obtained when a perfectly insulated calorimeter is calibrated. (b) These results might be obtained for a different calorimeter when the temperature readings are affected by heat loss and non-instantaneous heat transfer.

In Figure 3.4.3b, heat loss causes a negative slope to the line of the graph after the heater has been turned off. A more accurate estimate of the value of  $\Delta T$  can be found by **extrapolating** the line back to the time when heating commenced. The temperature change between this value and the initial temperature can be measured. A delay in the transfer of heat through the water can be observed by the continuing increase in temperature after the heater has been turned off. This can also be accounted for by the extrapolation method, although it is important to understand that the accuracy of this approach is limited.

### Worked example 3.4.1

#### CALCULATING THE CALIBRATION FACTOR OF A CALORIMETER BY ELECTRICAL CALIBRATION

A solution calorimeter was calibrated by passing 1.50 A through the electric heater for 50.5 s at a potential difference of 6.05 V. The temperature of the water in the calorimeter was initially 18.05°C and rose to 19.38°C during the calibration. Determine the calibration factor of the calorimeter.

Thinking	Working
Calculate the thermal energy released by the heater in the calorimeter when the electric current was passed through it. Use the equation: $E = VIt$	$E = VIt$ $= 6.05 \times 1.50 \times 50.5$ $= 458 \text{ J}$
Calculate the temperature change during the calibration. Use the equation: $\Delta T = \Delta T_{\text{final}} - \Delta T_{\text{initial}}$	$\Delta T = \Delta T_{\text{final}} - \Delta T_{\text{initial}}$ $= 19.38 - 18.05$ $= 1.33^\circ\text{C}$
Calculate the calibration factor by dividing the energy by the change in temperature. Use the equation: $CF = \frac{E}{\Delta T} = \frac{VIt}{\Delta T}$	$CF = \frac{E}{\Delta T}$ $= \frac{458}{1.33}$ $= 344 \text{ J } ^\circ\text{C}^{-1}$

### Worked example: Try yourself 3.4.1

#### CALCULATING THE CALIBRATION FACTOR OF A CALORIMETER BY ELECTRICAL CALIBRATION

A solution calorimeter was calibrated by passing 1.05 A through the electric heater for 120 s at a potential difference of 5.90 V. The temperature of the water in the calorimeter was initially 15.20°C and rose to 17.50°C during the calibration. Determine the calibration factor of the calorimeter.

If you know the calibration factor for a calorimeter, you can use it to determine the energy change that is responsible for a temperature change during a reaction in the calorimeter. These calculations are discussed later in this section.

#### Chemical calibration of solution calorimeters

A solution calorimeter may also be calibrated by performing a chemical reaction in the calorimeter that releases a known quantity of thermal energy and then measuring the resultant rise in temperature.

A highly soluble salt such as potassium nitrate ( $\text{KNO}_3$ ) could be used to calibrate a solution calorimeter. The **enthalpy of solution**,  $\Delta H$ , of potassium nitrate is known to be  $+34.9 \text{ kJ mol}^{-1}$ .

Calibration is achieved by dissolving a known amount (in moles) of potassium nitrate in the calorimeter. The change in temperature ( $\Delta T$ ) is measured and used to calculate the calibration factor, as shown in Worked example 3.4.2.

The energy released or absorbed during calibration is calculated using  $\Delta H$  for the reaction and the amount of reactant ( $n$ ):

$$E = n \times \Delta H$$

The calibration factor is then calculated using:

$$CF = \frac{E}{\Delta T}$$

### Worked example 3.4.2

#### CALCULATING THE CALIBRATION FACTOR OF A CALORIMETER BY CHEMICAL CALIBRATION

A solution calorimeter was calibrated by completely dissolving 15.0 g of potassium nitrate in 200 mL of water in a calorimeter. ( $M(\text{KNO}_3) = 101.3 \text{ g mol}^{-1}$ ;  $\Delta H = +34.9 \text{ kJ mol}^{-1}$ ). The temperature of the water in the calorimeter decreased from 20.5°C to 10.0°C during the calibration. Determine the calibration factor of the calorimeter.

Thinking	Working
Determine the amount, in mol, of potassium nitrate that dissolves. Use the equation: $n = \frac{m}{M}$	$n = \frac{m}{M}$ $= \frac{15.0}{101.3}$ $= 0.148 \text{ mol}$
Calculate the thermal energy absorbed from the water when the potassium nitrate dissolves in the calorimeter. Use the equation $E = n \times \Delta H$ , noting that the energy will be in kJ.	$E = n \times \Delta H$ $= 0.148 \times 34.9$ $= 5.17 \text{ kJ}$
Calculate the temperature change during the calibration. Use the equation: $\Delta T = \Delta T_{\text{final}} - \Delta T_{\text{initial}}$	$\Delta T = \Delta T_{\text{final}} - \Delta T_{\text{initial}}$ $= 20.5 - 10.0$ $= 10.5^\circ\text{C}$
Calculate the calibration factor, CF, by dividing the energy by the change in temperature. Use the equation: $CF = \frac{E}{\Delta T}$	$CF = \frac{E}{\Delta T}$ $= \frac{5.17}{10.5}$ $= 0.492 \text{ kJ } ^\circ\text{C}^{-1}$ $= 492 \text{ J } ^\circ\text{C}^{-1}$

### Worked example: Try yourself 3.4.2

#### CALCULATING THE CALIBRATION FACTOR OF A CALORIMETER BY CHEMICAL CALIBRATION

A solution calorimeter was calibrated by completely dissolving 8.20 g of potassium nitrate in 200 mL of water in a calorimeter. ( $M(\text{KNO}_3) = 101.3 \text{ g mol}^{-1}$ ;  $\Delta H = +34.9 \text{ kJ mol}^{-1}$ ). The temperature of the water in the calorimeter decreased from 20.0°C to 13.5°C during the calibration. Determine the calibration factor of the calorimeter.

## USING A CALIBRATED CALORIMETER TO DETERMINE THE ENTHALPY OF REACTION

A solution calorimeter may be calibrated by either the electrical or the chemical calibration methods described earlier in this section.

While it is inappropriate for a solution calorimeter to be used to determine the enthalpy of combustion for either a food or fuel, the enthalpy of reaction for many other reactions may be determined. These include acid–base reactions, reactions of a metal with an acid and dissolving a solid in water.

The calibration factor is used to determine the energy ( $E$ ) that is responsible for the temperature change that occurs during the reaction:

$$E = CF \times \Delta T$$



Note that the value of  $\Delta T$  is the change in temperature that occurs when the reaction is carried out in the calorimeter, not when the calorimeter is calibrated.

The enthalpy change,  $\Delta H$ , in  $\text{kJ mol}^{-1}$ , is calculated by dividing the energy change, in kJ, by the amount of the limiting reactant.

$$\Delta H = \frac{E}{n}$$

### Worked example 3.4.3

#### CALCULATING THE ENTHALPY OF SOLUTION USING A SOLUTION CALORIMETER

A solution calorimeter has a calibration factor of  $400 \text{ J } ^\circ\text{C}^{-1}$ . If the temperature decreases by  $1.375^\circ\text{C}$  when  $9.00 \text{ g}$  of glucose ( $M = 180.0 \text{ g mol}^{-1}$ ) is dissolved in water in the calibrated calorimeter, calculate the enthalpy of solution of glucose.

Thinking	Working
Calculate the amount of reactant in moles, using the equation: $n = \frac{m}{M}$	$n = \frac{m}{M}$ $= \frac{9.00}{180.0}$ $= 0.0500 \text{ mol}$
Calculate the heat energy released or absorbed, in kJ, using the equation: $E = CF \times \Delta T$	$E = CF \times \Delta T$ $= 400 \times 1.375$ $= 550 \text{ J}$ $= 0.550 \text{ kJ}$
Calculate the energy released or absorbed per mole. $\Delta H = \frac{E}{n}$	$\Delta H = \frac{E}{n}$ $= \frac{0.550}{0.0500}$ $= 11.0 \text{ kJ mol}^{-1}$
State the enthalpy of solution with the correct sign, remembering that a temperature increase indicates an exothermic process and a temperature decrease indicates an endothermic process.	$\Delta H = +11.0 \text{ kJ mol}^{-1}$ (3 significant figures)

### Worked example: Try yourself 3.4.3

#### CALCULATING THE ENTHALPY OF SOLUTION

A solution calorimeter has a calibration factor of  $396.4 \text{ J } ^\circ\text{C}^{-1}$ . If the temperature decreases by  $1.65^\circ\text{C}$  when  $41.587 \text{ g}$  of sucrose,  $\text{C}_{12}\text{H}_{22}\text{O}_{11}$  ( $M = 342.0 \text{ g mol}^{-1}$ ), is dissolved in water in the calibrated calorimeter, calculate the enthalpy of solution of sucrose.

## 3.4 Review



### SUMMARY

- Calorimetry is the experimental method by which the heat energy released or absorbed in a chemical reaction or physical process is measured.
- A calorimeter is an instrument that measures energy changes in a reaction. It is made up of an insulated container of water in which the reaction occurs, with a stirrer and thermometer to measure the temperature change during the reaction. A lid is an important part of the insulation.
- A solution calorimeter is an insulated container that holds a known volume of water and in which a reaction in solution, such as dissolution of a solid or a neutralisation reaction, can be carried out.
- Calorimeters can be calibrated to establish the relationship between the energy transferred to the water and the temperature change in the calorimeter. Calibration involves adding a known quantity of heat energy from an electrical source or from a chemical source.
- For electrical calibration of a calorimeter, the calibration factor is calculated using the equation  $CF = \frac{E}{\Delta T}$  where  $E = VIt$ . ( $V$  is the voltage, in volts,  $I$  is the current, in amps, and  $t$  is the time, in seconds.)
- During electrical calibration, the value of the temperature change,  $\Delta T$ , may be affected by heat loss from the calorimeter, so a temperature–time graph may be used to determine a more accurate value of  $\Delta T$ .
- When a reaction is carried out in a calibrated calorimeter, the energy released by the reaction is determined using the equation  $E = CF \times \Delta T$ .

### KEY QUESTIONS

#### Knowledge and understanding

- 1 Use the following terms to complete the statements about calorimeters:  
absorbed; accuracy; precision; decreases; endothermic; exothermic; increases; lid; lost; released; required; stays the same; thermometer  
Insulating a calorimeter improves the \_\_\_\_\_ of measurement of the quantity of energy \_\_\_\_\_ or absorbed by a chemical reaction.  
Heat energy can be \_\_\_\_\_ from a calorimeter, so a \_\_\_\_\_ is a useful form of insulation.  
If the reaction occurring in a calorimeter is exothermic, the temperature of the water \_\_\_\_\_. If the reaction occurring is \_\_\_\_\_ the temperature of the water in a calorimeter decreases.
- 2 A solution calorimeter was calibrated by completely dissolving 6.52 g of ammonium chloride in 200 mL of water in a calorimeter. ( $M(\text{NH}_4\text{Cl}) = 53.5 \text{ g mol}^{-1}$ ;  $\Delta H = +14.8 \text{ kJ mol}^{-1}$ .) The temperature of the water in the calorimeter decreased from 19.5°C to 14.3°C during the calibration. Determine the calibration factor of the calorimeter.
- 3 A solution calorimeter has a calibration factor of  $389 \text{ J } ^\circ\text{C}^{-1}$ . If the temperature decreases by 2.34°C when 9.60 g of citric acid,  $\text{C}_6\text{H}_8\text{O}_7$  ( $M = 192.0 \text{ g mol}^{-1}$ ), is dissolved in water in the calibrated calorimeter, calculate the enthalpy of solution of citric acid.
- 4 The enthalpy of solution of sodium hydrogen carbonate,  $\text{NaHCO}_3$  ( $M = 84.0 \text{ g mol}^{-1}$ ), can be determined using solution calorimetry. When 15.00 g of sodium hydrogen carbonate was dissolved in water in a calibrated calorimeter, the temperature decreased by 4.20°C. The calibration factor of the calorimeter was  $674.3 \text{ J } ^\circ\text{C}^{-1}$ . Calculate the enthalpy of solution of sodium hydrogen carbonate, including the sign to indicate whether it is an exothermic or endothermic reaction.

continued over page

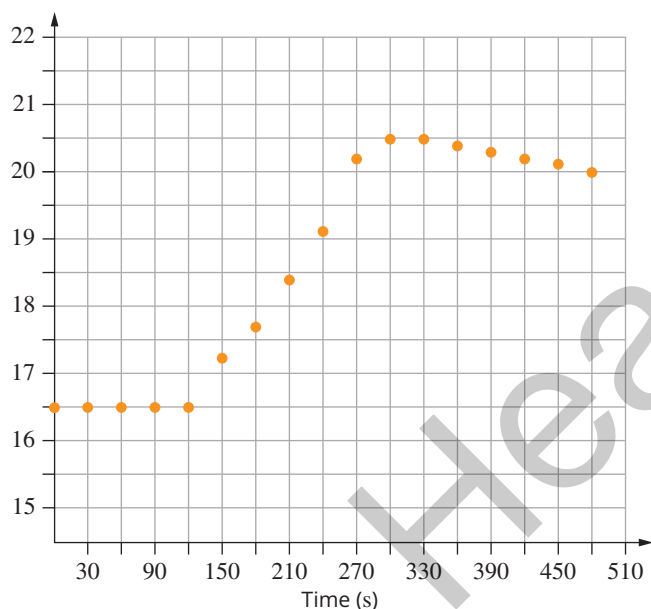
## 3.4 Review *continued*

### Analysis

**5** A simple polystyrene calorimeter was set up for an experiment. A 100 g mass of deionised water was used during the electrical calibration. The temperature of the calorimeter increased by  $9.5^{\circ}\text{C}$  when 3.8 A of current was passed through it at a potential of 6.0 V for 5 minutes.

- Determine the calibration factor for the calorimeter.
- What would be the effect on the calibration factor obtained from this experiment if 50 g of water is used instead of 100 g?

**6** A solution calorimeter was calibrated by the electrical method, using a current of 4.10 A for 150 s with a voltage of 5.75 V. The temperature–time graph for the calibration is shown in the figure below. The heater was turned on at  $t = 120$  s and turned off at  $t = 270$  s. Calculate the calibration factor for this calorimeter.



Graph of temperature against time during electrical calibration of a solution calorimeter

**7** Several mistakes or errors that could occur when calibrating a solution calorimeter are listed in the following table. For each mistake or error, indicate the likely effect on the calculated calibration factor by placing a tick in the appropriate column, and write an explanation for your answer in the final column.

Mistake or error	Calibration factor is too small	Calibration factor would be correct	Calibration factor is too large	Explanation for answer
The lid is left off during calibration.				
During electrical calibration, the voltmeter had a systematic error making it read too low.				
During chemical calibration, not all of the solid $\text{KNO}_3$ dissolved in the water of the calorimeter.				

## 3.5 Energy from fuels and food

Just as fossil fuels and biofuels supply you with energy for purposes such as heating, lighting and transport, food supplies the energy required for the millions of chemical reactions that occur in your body. You learnt in Chapter 2 that some fuels provide more energy per gram than others, and similarly, different foods provide different quantities of energy. When the energy contained in these fuels and food is transformed into forms of energy that we can use, much energy is lost in the transformation process.

In this section you will learn how to calculate the efficiency of energy transformations and you will learn about the different energy values of food.

### ENERGY TRANSFORMATIONS

When fuels are used or food is digested, the chemical energy is converted to a different form of energy. For example, a competitive cyclist may eat energy bars. Much of the chemical energy of the food is converted in the cyclist's body to mechanical and kinetic energy.

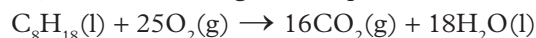
The conversion of the chemical energy of the food to the kinetic energy of the cyclist is an example of an energy transformation—energy is converted from one form to a different form. The use of fuels, including food, involves energy transformations. The chemical energy in a log on a campfire can be converted to thermal energy to heat a kettle of water over the fire.

When energy transformations occur, the total amount of energy is unchanged because energy cannot be created or destroyed. However, not all of the energy is converted to one specific form. In the case of the burning logs shown in Figure 3.5.1, very little of the chemical energy of the logs is converted to heat energy in the water.

### Energy transformation efficiency

The first law of thermodynamics tells us that energy is always conserved. However, when energy is being converted from one form into others, only some of the converted energy is likely to be useful. Scientists are interested in the efficiency of energy transformations. Many devices that we use every day can be described as **energy converters**. They transform energy from non-renewable or renewable sources into the form of energy that we require, such as electrical or mechanical energy.

The term **energy transformation efficiency** is used to describe the percentage of energy from a source that is converted to useful energy. For example, the engine in a petrol-fuelled car converts the chemical energy in petrol into thermal energy by combustion. You learned in Chapter 2 that the combustion of octane, the main component of petrol, occurs according to the equation:



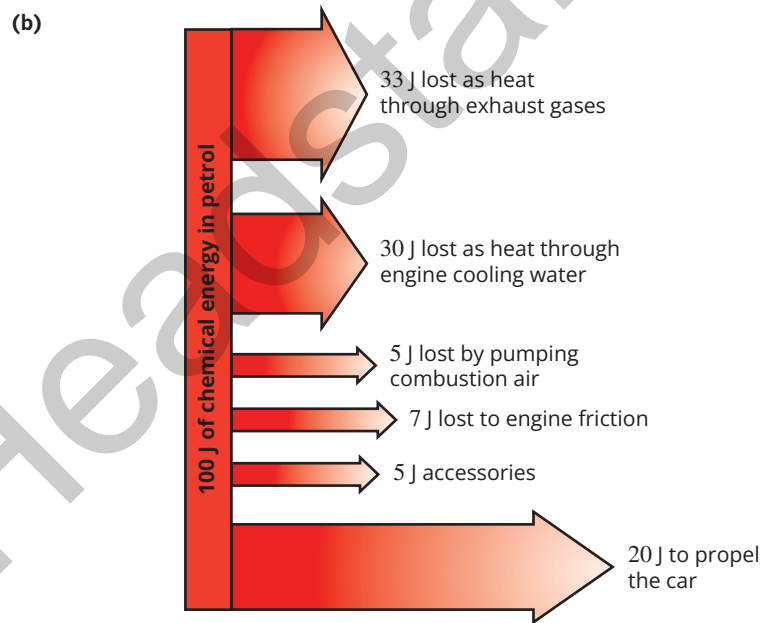
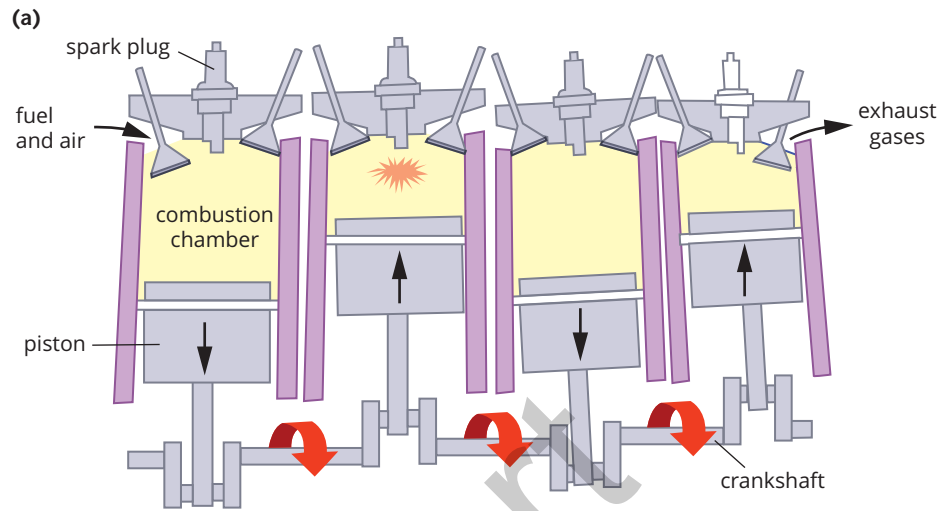
Some of the heat released by the reaction expands the gases in the engine's combustion chamber. As you can see in Figure 3.5.2a on the following page, these gases push the piston along the cylinder, causing the crankshaft to rotate and drive the wheels of the car.



**FIGURE 3.5.1** The chemical energy in the logs of the campfire is converted to thermal energy to boil the water in the kettle.



In Figure 3.5.2b the distribution of energy from the combustion of the petrol is illustrated. Only 20 J out of every 100 J is used to propel the car. The majority of the energy is transferred through the exhaust and the cooling system to the outside air as waste heat.



**i** Energy efficiency refers to the percentage of available energy that is transformed to the desired form of energy.

**TABLE 3.5.1** Energy efficiency of some energy converters

Device	Energy transformation	Typical efficiency
coal-fired power station	chemical to electrical	30%
lead-acid battery	chemical to electrical	70%
electric motor (as found in an electric car)	electrical to mechanical	77%
petrol engine	chemical to mechanical	20%
diesel engine	chemical to mechanical	30%
hydrogen fuel cell	chemical to electrical	60%

**FIGURE 3.5.2** Energy transformations in a petrol-fuelled car. (a) Combustion of the petrol generates heat and moves the pistons. (b) Only 20% of the energy released by the petrol is used to move the car.

The efficiency of an energy conversion can be viewed as an expression of the balance between energy input and useful energy produced.

$$\% \text{ energy efficiency} = \frac{\text{useful energy}}{\text{energy input}} \times \frac{100}{1}$$

The maximum value of **energy efficiency** is 100% but this value is not considered attainable in real situations.

In the case of the petrol-fuelled car described above, the energy efficiency is only 20%.

In comparison to the various devices described in Table 3.5.1, the energy efficiency of the human body is around 25%.

Throughout this book you will encounter further discussion of energy efficiencies as researchers are always striving to improve the sustainability of a process by increasing its efficiency.

In Section 3.3 you learnt how to determine the enthalpy of combustion of a fuel, such as an alcohol (see Figure 3.3.4, p XXX). When this experiment is carried out in a school laboratory, the results are rarely equal to the tabulated values due to heat loss to the surroundings. The energy efficiency of this experiment can be calculated using the theoretical value for the enthalpy of combustion and the calculated value from your results. This is illustrated in Worked example 3.5.1.

### Worked example 3.5.1

#### CALCULATING THE ENERGY TRANSFORMATION EFFICIENCY OF AN EXPERIMENT

A pair of students carried out an experiment to measure the enthalpy of combustion of methanol. They burned methanol in a spirit burner under a can of water and measured the temperature change of the water and the mass change of the spirit burner. From this experiment they determined the enthalpy of combustion of methanol to be  $-375 \text{ kJ mol}^{-1}$ .

The theoretical enthalpy of combustion of methanol was recorded in their data book as  $-726 \text{ kJ mol}^{-1}$ . Calculate the percentage efficiency of the energy transformation from chemical energy (of the methanol) to thermal energy (in the water).

Thinking	Working
Identify which energy best matches the description of useful energy and which is the energy input.	The 'useful energy' quantity is the energy that heated the water and enabled the calculation of the experimental value, $-375 \text{ kJ mol}^{-1}$ . The energy input quantity is the theoretical enthalpy of combustion: $-726 \text{ kJ mol}^{-1}$ .
Use the formula $\% \text{ energy efficiency} = \frac{\text{useful energy}}{\text{energy input}} \times \frac{100}{1}$ to calculate the percentage energy transformation efficiency.	$\begin{aligned} \% \text{ energy efficiency} &= \frac{\text{useful energy}}{\text{energy input}} \times \frac{100}{1} \\ &= \frac{-375}{-726} \times \frac{100}{1} \\ &= 51.7\% \end{aligned}$

### Worked example: Try yourself 3.5.1

#### CALCULATING THE ENERGY TRANSFORMATION EFFICIENCY OF AN EXPERIMENT

A pair of students carried out an experiment to measure the enthalpy of combustion of ethanol. They burned ethanol in a spirit burner under a can of water and measured the temperature change of the water and the mass change of the spirit burner. From this experiment they determined the enthalpy of combustion of ethanol to be  $-820 \text{ kJ mol}^{-1}$ .

The theoretical enthalpy of combustion of ethanol was recorded in their data book as  $-1360 \text{ kJ mol}^{-1}$ . Calculate the percentage efficiency of the energy transformation from chemical energy (of the ethanol) to thermal energy (in the water).

In some cases, you will know the percentage efficiency of a process. Worked example 3.5.2 shows you how to include this in your calculations.

### Worked example 3.5.2

#### ALLOWING FOR % EFFICIENCY IN ENERGY TRANSFORMATIONS

A propane burner is used to heat a saucepan containing water. The % efficiency of this process is 35.0%. Calculate the energy transferred to the water from the complete combustion of 0.320 mol of propane. The heat of combustion is 2220 kJ mol<sup>-1</sup>.

Thinking	Working
Calculate the amount, in mol, of the fuel.	0.320 mol (given in the question)
Calculate the quantity of energy released (the energy input) by the combustion of the fuel.	$\begin{aligned} \text{Energy released} &= n \times \Delta H \\ &= 0.320 \times 2220 \\ &= 710.4 \text{ kJ} \end{aligned}$
Rearrange the formula $\% \text{ energy efficiency} = \frac{\text{useful energy}}{\text{energy input}} \times \frac{100}{1}$ to Useful energy = energy input × % energy efficiency so that the energy available to heat the water can be determined.	$\begin{aligned} \text{Useful energy} &= \text{energy input} \times \% \text{ energy efficiency} \\ &= 710 \times \frac{35}{100} \\ &= 249 \text{ kJ} \end{aligned}$

### Worked example: Try yourself 3.5.2

#### ALLOWING FOR % EFFICIENCY IN ENERGY TRANSFORMATIONS

The % efficiency of octane (petrol) in a particular car model is 42.0%. Calculate the energy released from the complete combustion of 22.0 g of octane. The heat of combustion of octane is 5460 kJ mol<sup>-1</sup>.

In other cases, you will know the percentage efficiency of a process and the energy that was transferred to the water, but you will be asked to determine the theoretical heat of combustion of the fuel. Worked Example 3.5.3 shows you how to include this in your calculations.

### Worked example 3.5.3

#### CALCULATING HEAT OF COMBUSTION USING % EFFICIENCY IN ENERGY TRANSFORMATIONS

A student burnt 0.0250 mol of propan-2-ol in a spirit burner to heat 200 g of water in a can. The % efficiency of this process was 65.0%. During the experiment, the temperature of the water in the can increased by 39.0°C. Calculate the heat of combustion of propan-2-ol. (Remember that the specific heat capacity of water is 4.18 J °C<sup>-1</sup> g<sup>-1</sup>.)

Thinking	Working
Calculate the quantity of energy that heated the water—this is the useful energy. Use the formula $q = mc\Delta T$ to calculate the useful energy, because the can does not have a calibration factor.	$q = mc\Delta T$ $= 200 \times 4.18 \times 39.0$ $= 32\,604 \text{ J}$ $= 32.604 \text{ kJ}$
Convert the % energy efficiency to a decimal to assist with calculations. Divide % energy efficiency by 100.	$\frac{\% \text{ energy efficiency}}{100}$ $= \frac{65.0}{100}$ $= 0.650$
Rearrange the formula $\% \text{ energy efficiency} = \frac{\text{useful energy}}{\text{energy input}} \times \frac{100}{1}$ to energy input = $\frac{\text{useful energy}}{\% \text{ energy efficiency (as a decimal)}}$ so that the energy released by the fuel can be determined.	$\text{energy input} = \frac{32.604}{0.650}$ $= 50.16 \text{ kJ}$
Remembering that the energy released is represented by the energy input, rearrange the formula: energy released = $n \times \Delta H$ to: $\Delta H = \frac{\text{energy released}}{n}$	$\Delta H = \frac{50.16}{0.0250}$ $= 2006$ $= 2.01 \times 10^3 \text{ kJ mol}^{-1}$ The heat of combustion of propan-2-ol is 2.01 × 10 <sup>3</sup> kJ mol <sup>-1</sup>

### Worked example: Try yourself 3.5.3

#### CALCULATING HEAT OF COMBUSTION USING % EFFICIENCY IN ENERGY TRANSFORMATIONS

A student is burning butan-1-ol in a spirit burner to heat 200 g of water in a can. The % efficiency of this process is 57.0%. During the combustion of 0.0105 mol of butan-1-ol, the temperature of the water in the can increased by 19.2°C. Calculate the heat of combustion of butan-1-ol. (Remember that the specific heat capacity of water is 4.18 J °C<sup>-1</sup> g<sup>-1</sup>.)



## CHEMFILE

### The importance of cellulose

Even though humans are unable to digest and obtain the energy available from cellulose, a diet high in cellulose provides 'bulk' to aid the passage of food through the digestive system. Such 'bulk' helps prevent constipation and reduce the risk of bowel cancer, also known as colorectal cancer. Bowel cancer is the second most common cancer in men and women in Australia and is more common in people over the age of 50. High red meat consumption (in excess of 500 g per week), especially processed meats, increases the risk of developing bowel cancer. This risk can be reduced by eating a healthy diet with plenty of fresh fruit and vegetables as shown in the figure below.



A diet high in fruit and vegetables reduces the risk of bowel cancer.

**i** The energy content of the three major nutrients is:

carbohydrates	16 kJ g <sup>-1</sup>
fats	37 kJ g <sup>-1</sup>
proteins	17 kJ g <sup>-1</sup>

## ENERGY AVAILABLE TO THE BODY

Food supplies your body with **nutrients** that are needed to keep you alive. These nutrients are large **biomolecules** that are used by the body to provide energy, regulate growth, and maintain and repair body tissue. The three main nutrient classes are carbohydrates, proteins and fats. The **energy value** or energy content of these nutrients was discussed in Section 2.5. The energy released when food is burned is often greater than the energy that is available for the human body to use after the food has been digested. This can be due to:

- incomplete absorption of nutrients by the body after digestion of the food
- incomplete oxidation of nutrients, such as proteins and insoluble fibre
- heat loss; not all of the energy released by the digestion of food is available for use in cells as some is lost as waste heat.

Even within a particular nutrient class, such as the carbohydrates, the energy available to the body can vary. Starch is readily digested by humans and is the primary source of glucose, so most of the energy from starch is obtained during digestion. On the other hand, humans lack the enzyme cellulase, which is required to break down cellulose so that energy can be obtained. Cellulose in foods such as vegetables is often referred to as **dietary fibre** and provides no energy.

A balanced diet is made up of a variety of foods containing carbohydrates, proteins and fats. Each of these three major nutrients provides a different quantity of energy per gram. In Section 2.3 you learnt that the energy available to the body from carbohydrates is 16 kJ g<sup>-1</sup>, from fats and oils is 37 kJ g<sup>-1</sup> and from proteins is 17 kJ g<sup>-1</sup>.

You can use these values to calculate the energy value of different foods.

## CALCULATING THE ENERGY VALUE OF FOODS

While you may distinguish between different foods based on how delicious they are to you, a more scientific difference is the variation in percentages of carbohydrates, fat and protein in each food. In Figure 3.5.3 you can see that foods with a high percentage of carbohydrates and fat also have a high energy content.

Since it is important to know how much energy is available from particular foods, processed foods are labelled with their energy content, as well as percentages of carbohydrates, fats and protein.



Food	Carbohydrate (g/100g)	Fat (g/100g)	Protein (g/100g)	Energy content (kJ/100g)
milk chocolate	52	30	8	2130
donut	51	23	5	1803
apple	14	0.2	0.3	247

**FIGURE 3.5.3** Foods that have a high percentage of carbohydrate and fat have a high energy content.

When the percentage composition of a food is known, its energy value can be calculated using the percentages and the energy available to the body for each nutrient. The energy value of foods may be expressed in kJ/100g or kJ g<sup>-1</sup>. The composition and energy values of a range of foods are given in Table 3.5.2.

**TABLE 3.5.2** The composition and energy value of a range of foods

Food	Carbohydrate (%)	Protein (%)	Fats and oils (%)	Energy value (kJ g <sup>-1</sup> )
white rice	79	7	negligible	15.2
wholemeal bread	39	11	4	9.7
avocados	6	2	17	7.2
roasted peanuts	18	26	50	24.3
pizza (with cheese)	31	11	8	10.0
apples	53	4	8	11.7
almonds	19	19	54	25.4

### Worked example 3.5.4

#### CALCULATING THE ENERGY VALUE OF FOODS

Labelling on a sample of unsalted cashews indicates they contain 29.0% carbohydrates, 18.0% protein and 46.0% fat. The remaining 7.0% is water, which does not supply energy.

Calculate the energy value of the cashews, in kJ g<sup>-1</sup>.

Thinking	Working
Use the information on page xxx to determine the available energy for each nutrient type.	carbohydrate: 16 kJ g <sup>-1</sup> protein: 17 kJ g <sup>-1</sup> fat: 37 kJ g <sup>-1</sup>
Assuming that there is 100 g of the sample, multiply each percentage of the nutrient by the available energy per gram for the nutrient type.	carbohydrate: 29.0 g × 16 kJ g <sup>-1</sup> = 464 kJ protein: 18.0 g × 17 kJ g <sup>-1</sup> = 306 kJ fat: 46.0 g × 37 kJ g <sup>-1</sup> = 1702 kJ
Find the sum of the energies for the three nutrient types and divide by 100 to find the energy value in kJ g <sup>-1</sup> .	energy value = $\frac{464 + 306 + 1702}{100}$ = 24.7 kJ g <sup>-1</sup>

### Worked example: Try yourself 3.5.4

#### CALCULATING THE ENERGY VALUE OF FOODS

Labelling on a sample of white bread indicates it contains 53.0% carbohydrates, 8.0% protein and 4.0% fat. The remaining 35.0% is water, which does not supply energy.

Calculate the energy value of the bread, in kJ g<sup>-1</sup>.

## 3.5 Review



### SUMMARY

- Energy efficiency refers to the percentage of available energy that is transformed to the desired form of energy as calculated by:  
$$\% \text{ energy efficiency} = \frac{\text{useful energy}}{\text{energy input}} \times \frac{100}{1}$$
- The energy value of fuels and foods is measured in  $\text{kJ g}^{-1}$ .
- The energy value of a food can be calculated using data for the available energy of the various components of the food and the percentages of each component of the food.

### KEY QUESTIONS

#### Knowledge and understanding

- 1 Calculate the mass of each of the following nutrients that supplies the same energy as that released by 10.0 g of carbohydrates.
  - a protein
  - b fat
- 2 The use of coal to produce electrical energy requires several energy transformations, as shown in Figure 3.1.5 on p. XX.
  - a What is an energy transformation?
  - b How do energy transformations impact on the usefulness of a fuel?
- 3 An 18.0 g chocolate frog contains 5.3 g fat, 11.0 g of carbohydrates and 1.4 g protein.
  - a Calculate the total energy in the chocolate frog.
  - b Calculate the percentage of total energy that is a consequence of the fat content.

#### Analysis

- 4
  - a Butanol has a heat of combustion of  $2670 \text{ kJ mol}^{-1}$ . Calculate the mass, in kg, of butanol that is required to produce 200 MJ of energy.
  - b A 375 g sample of butanol is used to heat a drum of water. The efficiency of this process is 38.0%. Calculate the quantity of energy transferred to the water.
- 5 Labelling on a jar of pesto indicates it contains 3.3% carbohydrates, 3.9% protein and 21.4% fat. The remaining mass is water. Calculate the energy value of the pesto, in  $\text{kJ g}^{-1}$ .

- 6 A 15.0 g sample of butter is reported as having 425 kJ of energy. If all of this energy came from fat sources in the butter, calculate the percentage by mass of fat in the butter sample.
- 7 A student conducted a practical investigation to experimentally determine the heat of combustion of propan-1-ol using the apparatus shown in Figure 3.3.3. The following data is obtained:

mass of spirit burner before burning	3.565 g
mass of spirit burner after burning	3.284 g
mass of water	100 g
initial temperature of water	18.6 °C
final temperature of water	30.8 °C

- a Determine the quantity of energy, in kJ, that was used to heat the water.
  - b Calculate the heat of combustion, in  $\text{kJ mol}^{-1}$ , of propan-1-ol according to this data.
  - c The heat of combustion of propan-1-ol obtained from a secondary source is  $2021 \text{ kJ mol}^{-1}$ . What was the percentage energy efficiency of this calculation of the heat of combustion of propan-1-ol?
  - d Describe three ways the student could improve this experiment to increase the accuracy of the result.
- 8 A student burnt 0.0325 mol of methanol in a spirit burner under 250 g of water in a can. The temperature of the water increased by  $16.9^\circ\text{C}$  during the experiment. If the % efficiency of this process was 75.0%, calculate the heat of combustion of methanol.

# Chapter review



# 03

## KEY TERMS

amount	energy transformation	mass–volume
biomolecule	efficiency	stoichiometry
bomb calorimeter	energy value	molar volume
calibrated	enthalpy of solution	mole ratio
calibration factor	excess reactant	nutrient
calorimeter	exothermic	reagent
calorimetry	extrapolating	solution calorimeter
coefficients	greenhouse gas	specific heat capacity
complete combustion	heat of combustion	standard laboratory
dietary fibre	in excess	conditions, SLC
endothermic	incomplete combustion	stoichiometry
energy content	limiting reactant	volume–volume
energy converter	mass–mass	stoichiometry
energy efficiency	stoichiometry	
energy transformation		

## REVIEW QUESTIONS

### Knowledge and understanding

- Which one of the following nutrients will produce the most energy when consumed?  
A 15 g protein  
B 8 g fat  
C 20 g carbohydrate  
D 5 g fat and 10 g protein
- The enthalpy of combustion of propan-1-ol is determined experimentally to be  $-1800 \text{ kJ mol}^{-1}$ , while the theoretical value is  $-2018 \text{ kJ mol}^{-1}$ . Which of the following is the correct value of the percentage energy transformation efficiency in this experiment?  
A 10.8%  
B 12.1%  
C 50.0%  
D 89.2%
- Propane ( $\text{C}_3\text{H}_8$ ) burns in oxygen according to the equation:  
$$\text{C}_3\text{H}_8(\text{g}) + 5\text{O}_2(\text{g}) \rightarrow 3\text{CO}_2(\text{g}) + 4\text{H}_2\text{O}(\text{g})$$
  
6.70 g of propane was burned in excess oxygen.
  - What mass of carbon dioxide would be produced?
  - What mass of oxygen would be consumed in the reaction?
  - What mass of water would be produced?
- 5.0 g of magnesium burns completely in 20 g of oxygen gas according to the equation:  
$$2\text{Mg}(\text{s}) + \text{O}_2(\text{g}) \rightarrow 2\text{MgO}(\text{s})$$
  
What mass of oxygen gas is in excess?
- Large quantities of coal are burned in Australia to generate electricity, in the process generating significant amounts of the greenhouse gas carbon dioxide. The equation for this combustion reaction can be written as:  
$$\text{C}(\text{s}) + \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g})$$
  
Determine the mass of carbon dioxide produced by the combustion of 1.0 tonne ( $10^6 \text{ g}$ ) of coal, assuming that the coal is pure carbon.
- Octane is one of the main constituents of petrol. It burns according to the equation:  
$$2\text{C}_8\text{H}_{18}(\text{g}) + 25\text{O}_2(\text{g}) \rightarrow 16\text{CO}_2(\text{g}) + 18\text{H}_2\text{O}(\text{g})$$
  
Calculate the volume of octane used if 50.0 L of carbon dioxide was produced, and all gas volumes were measured at SLC.
- Calculate the quantity of energy that has been transferred to 50.0 g of water in a test tube when the temperature of the water increases from  $18.5^\circ\text{C}$  to  $29.8^\circ\text{C}$ .
- Calculate the quantity of energy released when 806 g of butan-1-ol reacts with excess air according to the equation:  
$$\text{C}_4\text{H}_9\text{OH}(\text{g}) + 6\text{O}_2(\text{g}) \rightarrow 4\text{CO}_2(\text{g}) + 5\text{H}_2\text{O}(\text{l})$$
  
$$\Delta H = -2677 \text{ kJ}$$
- Propane ( $\text{C}_3\text{H}_8$ ) burns in oxygen according to the equation:  
$$\text{C}_3\text{H}_8(\text{g}) + 5\text{O}_2(\text{g}) \rightarrow 3\text{CO}_2(\text{g}) + 4\text{H}_2\text{O}(\text{g})$$
  
300 g of propane reacts with 2.00 kg of oxygen.
  - Determine which reactant is in excess and the mass by which it is in excess,
  - Calculate the volume of carbon dioxide produced, at SLC.
  - Calculate the volume of carbon dioxide produced if the conditions were  $15^\circ\text{C}$  and 150 kPa.

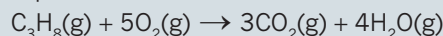
- 10** The equation  $q = mc\Delta T$  describes the relationship between heat energy and the temperature change of a substance being heated. It is used to determine how much energy has been absorbed by a material, such as a mass of water, when its temperature increases by a measured amount.
- Which one of the following statements about this equation is correct?
- A** The specific heat capacity,  $c$ , depends on the material that is generating the heat energy.  
**B** The mass,  $m$ , represents the mass of the material that is being heated.  
**C**  $q$  is measured in  $\text{kJ } ^\circ\text{C}^{-1}$ .  
**D**  $\Delta T$  refers to the change in temperature of the fuel generating the heat.
- 11** Which one of the following is most likely to be the temperature change in 100.0 mL of water in a beaker with no insulation and no lid, when 2.10 kJ of energy is produced by a burning biscuit under the beaker?
- A**  $3.20^\circ\text{C}$   
**B**  $5.02^\circ\text{C}$   
**C**  $5.50^\circ\text{C}$   
**D**  $6.30^\circ\text{C}$
- 12** Pressed peat has a lower energy content than coal, and produces more carbon dioxide, and less smoke. It has a lower sulfur content, and is considered renewable. A sample of pressed peat has an energy content of  $12.8 \text{ kJ g}^{-1}$ . Calculate the mass of water, in kg, that could be heated from  $20.0^\circ\text{C}$  to  $60.0^\circ\text{C}$  by the transfer of all of the energy from the combustion of a sample of peat with a mass of 17.0 g.
- 13** When 1.01 g of solid potassium nitrate was dissolved in a calorimeter, the temperature dropped by  $0.672^\circ\text{C}$ . When the calorimeter was calibrated, a current of 1.50 A applied at a potential difference of 5.90 V for 60.0 s caused a temperature rise of  $0.456^\circ\text{C}$ .
- a** Write an equation for the dissolution of potassium nitrate.  
**b** Determine the calibration factor of the calorimeter.  
**c** Calculate the energy change during the reaction.  
**d** Calculate the enthalpy change of reaction for the equation written in part **a**.
- 14** Labelling for a stuffed pepper indicates it contains 16% carbohydrates, 13% protein and 6.0% fat. The rest of the pepper is water.
- Calculate the energy value of the stuffed pepper, in  $\text{kJ g}^{-1}$ .

### Application and analysis

- 15** A calorimeter is calibrated by the electrical calibration method. The temperature rise was measured as the difference between the highest temperature reached and the initial temperature.
- Describe the most likely effect on the calibration factor if the quantity of current passing through the calorimeter was higher than measured.
- 16** Energy bars are often popular with runners and cyclists due to their portability and high energy derived largely from carbohydrates and proteins. One particular energy bar contains 4.0 g fat, 44.0 g carbohydrates and 10.0 g of protein.
- a** Calculate the total energy in one energy bar.  
**b** An athlete running with moderate intensity will consume approximately 2400 kJ per hour of running. Calculate how many energy bars they would need consume to sustain their energy needs if they were going for a 1.5 hour-long run.
- 17** Methane burns in excess oxygen according to the equation:
- $$\text{CH}_4(\text{g}) + 2\text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{g})$$
- This reaction produces 5.0 L of carbon dioxide at SLC. Assuming all volumes are measured at the same temperature and pressure, calculate the:
- a** volume of methane used  
**b** volume of oxygen used  
**c** mass of greenhouse gases produced.
- 18** Octane ( $\text{C}_8\text{H}_{18}$ ), the principal component of petrol, undergoes combustion with a plentiful supply of oxygen according to the equation:
- $$2\text{C}_8\text{H}_{18}(\text{l}) + 25\text{O}_2(\text{g}) \rightarrow 16\text{CO}_2(\text{g}) + 18\text{H}_2\text{O}(\text{l})$$
- $$\Delta H = -10\,900 \text{ kJ}$$
- a** What mass of octane must be burnt in order to yield 100 kJ of energy?  
**b** Given that density is calculated as  $\frac{\text{mass}}{\text{volume}}$  and the density of octane is  $0.698 \text{ g mL}^{-1}$ , determine the energy released when 50.0 L of liquid octane undergoes complete combustion.
- 19** An indoor gas heater burns propane ( $\text{C}_3\text{H}_8$ ) at a rate of 12.7 g per minute. Calculate the minimum mass of oxygen per minute, in g, that needs to be available for the complete combustion of propane.

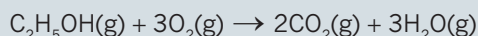


- 20 Propane ( $C_3H_8$ ) burns completely in oxygen according to the equation:



- a When 80 mL propane and 500 mL of oxygen are reacted, both at SLC, which is in excess and by what volume?
- b What is the total volume of greenhouse gases produced in the reaction?
- c What change in the total volume of all of the gases has occurred as a result of this reaction?
- d The amount of energy released per tonne of carbon dioxide produced is determined to be  $1.68 \times 10^4 \text{ MJ t}^{-1}$ . Determine the  $\Delta H$  value, in  $\text{kJ mol}^{-1}$ , for the thermochemical equation. (1 tonne =  $10^6 \text{ g}$ ). Show your working.

- 21 In a sealed reaction to produce energy from the combustion of ethanol, 80.0 L of carbon dioxide was collected at SLC. The reaction is

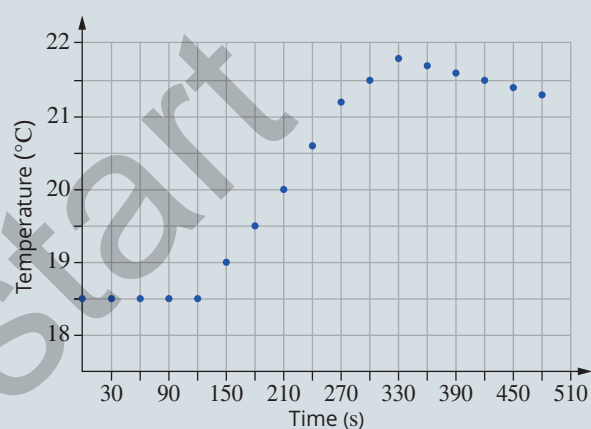


If 10.0 g of ethanol was in the sealed container when it was ignited, was the ethanol limiting or in excess? Explain your answer.

- 22 A peanut with a mass of 1.340 g was burned under a steel can containing 200 g of water. After the flame went out, 0.640 g of the peanut remained and the temperature of the water had risen by  $17.3^\circ\text{C}$ . Calculate the energy content of the peanut in  $\text{kJ g}^{-1}$ , assuming all the energy from the burning nut is transferred to the water.
- 23 0.254 g of black coal was burnt and used to heat 300 g of water. The temperature of the water rose from  $18.25^\circ\text{C}$  to  $24.92^\circ\text{C}$ . Calculate the heat of combustion, in  $\text{kJ g}^{-1}$ , of the coal. Assume all of the heat released during combustion was used to heat the water.
- 24 A temperature rise of  $5.52^\circ\text{C}$  was observed when 0.0450 g of ethane ( $C_2H_6$ ) was burnt and used to heat 100.0 g of water initially at  $25^\circ\text{C}$ . Use this information to write a balanced thermochemical equation for the complete combustion of ethane.
- 25 A 500 g mass of water in a beaker was heated using the energy released by the combustion of a 3.00 g piece of wood. The observed temperature rise was  $22.9^\circ\text{C}$ .
- a Calculate the heat of combustion of wood in  $\text{kJ g}^{-1}$ .
- b Would the heat of combustion you calculated in part a be higher or lower than the actual heat of combustion of the wood? Give reasons for your answer.

- 26 A solution calorimeter was calibrated electrically. A potential of 9.80 V was applied for 270 s with a current of 3.60 A and the temperature increased by  $3.49^\circ\text{C}$ . 3.00 g of sucrose (molar mass  $342.0 \text{ g mol}^{-1}$ ) was dissolved completely in the same calibrated solution calorimeter. During the dissolving of the sucrose, the temperature of the water decreased by  $18.14^\circ\text{C}$ . What is the heat of solution of sucrose in  $\text{kJ mol}^{-1}$ ?

- 27 The following figure shows the temperature–time graph plotted by the data-logging equipment used by a student to determine the change in temperature during the calibration of a solution calorimeter. The heater was turned on at  $t = 120 \text{ s}$  and turned off at  $t = 300 \text{ s}$ .



Temperature–time graph used for calibration of a calorimeter

The student ignored the graph and instead used the formula  $\Delta T = T_{\text{final}} - T_{\text{initial}}$  to determine  $\Delta T$  and hence the calibration factor.

The student then correctly made the additional measurements needed to determine the enthalpy change of a reaction.

Describe the effect of *not* compensating for the heat loss on the following.

- a the value of  $\Delta T$
- b the calibration factor calculated for the calorimeter
- c the value of the enthalpy change measured by the calorimeter

## CHAPTER REVIEW CONTINUED

- 28** A student is having difficulty with the calculations for the calibration of a solution calorimeter and the calculations that follow. The student wants to determine the heat energy per gram required to dissolve a packet serving of jelly crystals in water. The student's data for the calibration are shown in Tables 3.6.1 and 3.6.2.

**TABLE 3.6.1** Calibration data

Volume of water in calorimeter (mL)	250
Voltage (V)	6.30
Current (amps)	2.40
Time (s)	210
Initial temperature (°C)	17.5
Final temperature (°C)	20.3

**TABLE 3.6.2** Data from experiment in which jelly crystals are dissolved in water

Volume of water in calorimeter (mL)	250
Mass of jelly crystals (g)	85.0
Initial temperature (°C)	20.3
Final temperature (°C)	19.8

The student's calculations are as follows:

$$CF = 6.30 \times 2.40 \times 210 \times 2.8 \\ = 8890 \text{ J}$$

Heat energy per gram to dissolve the jelly crystals:

$$E = CF \times \Delta T \\ = 8890 \times 0.5 \\ = 4445 \text{ J}$$

$$\text{energy per gram} = \frac{85.0}{4445} \\ = 0.019 \text{ J g}^{-1}$$

0.019 J of energy is given out when 1 gram of jelly crystals dissolves. This is an endothermic reaction. Describe any errors that the student has made in the calculations, and then complete the calculation correctly.

