

PEARSON  
**Science**

STUDENT BOOK | WESTERN AUSTRALIA

9



# Chemical reactions: Rearranging atoms

Chemicals react together in specific ways to produce new substances that can be represented in a chemical equation. A word equation is a simple description of the reaction with reactants on the left side of the arrow and the products on the right. The law of conservation of mass states that mass cannot be created or destroyed in a chemical reaction and any new product is formed by the rearrangement of the atoms present in the reactants.

While mass is conserved in a reaction, energy can be absorbed or released, resulting in a change of temperature. Exothermic reactions, such as combustion, release energy and cause a rise in temperature, whereas endothermic reactions absorb energy and result in a reduction in temperature.

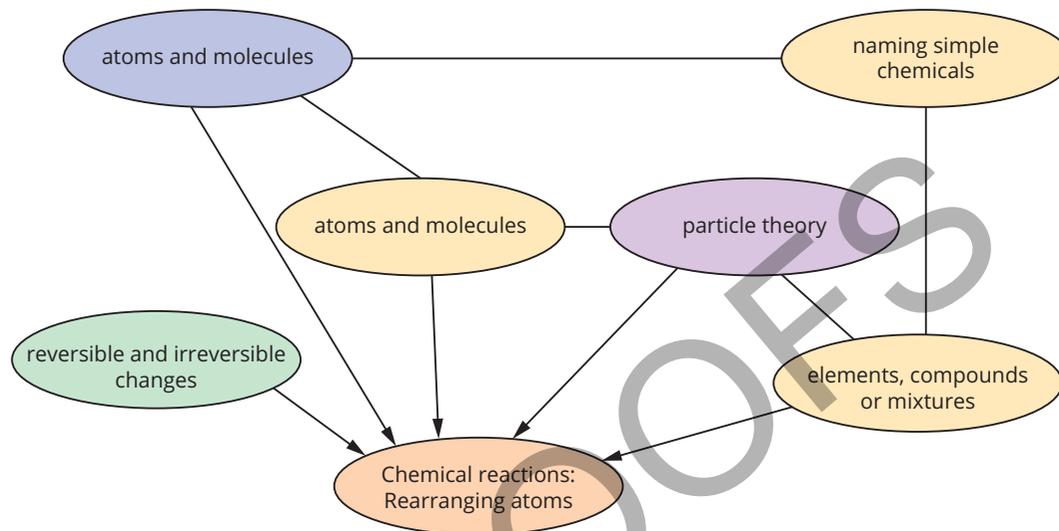
In this topic, you will learn about classifying reactions to predict the formation of products. You will also learn how to balance chemical equations to ensure that they follow the law of conservation of mass.

## Learning intentions

- To understand the key components of chemical reactions **xx**
- To be able to measure and record accurate data to provide evidence for the conservation of mass in a chemical reaction **xx**
- To understand that similarities exist between different types of chemical reactions **xx**
- To be able to carry out, describe, represent and compare a range of chemical reactions **xx**
- To understand that atomic rearrangement occurs when reactants are converted into products during a chemical reaction **xx**
- To understand how to write balanced equations for simple chemical reactions **xx**
- To be able to evaluate why green chemistry processes are being adopted by manufacturers **xx**
- To be able to observe combustion reactions and model the atomic rearrangements occurring during the reactions **xx**

# Chemical reactions: Rearranging atoms

The key concepts that you will use in this topic:



The following prior knowledge questions will help to support your learning in the topic and can be attempted before the first lesson.

## Chemical and physical changes

**1** Zia wants to make toffee. She warms water on a gas stove. Next, she adds sugar and stirs until the sugar is fully dissolved. She boils the sugary water. Unfortunately, Zia got distracted. All the water evaporated and the sugar burnt.

Categorise the following processes as chemical or physical change.

- a gas burning on the gas stove
- b dissolving sugar in water
- c boiling water
- d evaporating water
- e burning sugar

## Reversible and irreversible processes

**2** Identify whether the following processes as reversible or irreversible.

- a a glow stick
- b fireworks
- c freezing ice-cubes

## Elements, compounds and mixtures

- 3 The chemical formula for a commonly used black dye is  $C_{69}H_{117}N_9O_6RuS_3$ . Using a periodic table, list the elements in the black dye molecule.

## Exothermic and endothermic reactions

- 4 Both instant cold packs and heat packs are used for sports injuries. Chemical reactions form inside the pack, causing the surroundings to change temperature. Classify each type of pack as an exothermic reaction or an endothermic reaction.

## Word and chemical equations

- 5 Carbonyl diamide (urea) is the substance commonly used in industry, agriculture and medicine. It is produced by reacting carbon dioxide with ammonia. Water was also a product. Write out the word equation for the production of carbonyl diamide.

PAGE PROOFS

# 4.1 Chemical reactions

## Lesson overview

Chemical reactions are taking place all around you. Respiration, cooking food, burning fuels, and even operating an electric car: all these processes involve important chemical reactions. In a chemical reaction, new substances are made, but no new atoms are formed. They are simply rearranged or joined together differently.

In this lesson, you will learn about how the atoms inside the reactants (the substances initially present in a chemical reaction) get rearranged to form products (the substances produced in a chemical reaction) and how this process can be conveniently described using chemical equations.

**SC 1** I can identify the reactants and products in a chemical reaction

When a chemical reaction occurs, new substances are produced. The substances that react together are called **reactants**. The new substances produced are called **products**.

In a corrosion reaction that results in rusting, for example, iron, water, and oxygen are the reactants. The product formed is hydrated iron(III) oxide, otherwise known as rust (Figure 4.1.1).



**FIGURE 4.1.1** The iron in these railway tracks has reacted with oxygen and water to form rust.

Another type of a chemical reaction is combustion, in which fuels burn in oxygen. For example, when methane burns in oxygen, methane and oxygen are the reactants. Carbon dioxide and water are the products in addition to heat and light energy (Figure 4.1.2).

## Word and symbol equations

Chemical reactions are not usually described in full sentences, but in a shorter form called a **chemical equation**.

### Learning intention

To understand the key components of chemical reactions

### Success criteria

**SC 1:** I can identify the reactants and products in a chemical reaction.

**SC 2:** I can recall that mass is conserved during a chemical reaction.

**SC 3:** I can identify the names or formulas of reactants and products in familiar chemical reactions.

### KEY TERMS

**reactant** a substance that takes part in a chemical reaction

**product** a substance produced by a chemical reaction

**chemical equation** a representation of a chemical reaction that uses symbols and formulas to represent substances



**FIGURE 4.1.2** Methane gas burns in a gas hob, reacting with oxygen to produce carbon dioxide and water, as well as heat and light.

A chemical equation is always written in the following format, where reactants appear to the left of the arrow and products to the right:

reactants  $\rightarrow$  products

The arrow that separates the reactants from the products means ‘yields’ or ‘produces’. The single arrow tells you that the reaction is irreversible, that is, the products cannot turn back into reactants.

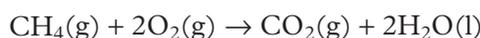
reactants  $\rightleftharpoons$  products

The arrow  $\rightleftharpoons$  indicates the chemical reaction is reversible; that is, the products can turn back into the reactants.

A **word equation** uses the names of the reactants and products. For example:

methane + oxygen  $\rightarrow$  carbon dioxide + water

A **balanced equation** uses chemical symbols and formulas to represent the reactants and products. For example:



In a balanced equation, the numbers of each atom on the left side of the arrow must be the same as on the right.

### KEY TERMS

**word equation** a chemical equation in which the reactants and products are identified by their chemical names

**balanced equation** a chemical equation that has the same number of each type of atom on both sides of the equation

### HINTS

The letters in the brackets represent the state (solid, liquid, gas, dissolved) which the substance exists in for the reaction. The letters represent:

solid

liquid

gas

aqueous (the substance is dissolved in water; *aqua* is the Latin word for water)

### SC 1 CHECK YOUR UNDERSTANDING

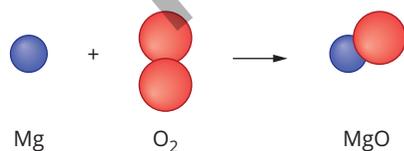
The chemical used in the food industry to provide orange flavour to food is called ethyl butanoate. It is formed by the reaction of ethanol with butanoic acid, and water is another product. Write out the word equation for this reaction.

### SC 2 I can recall that mass is conserved during a chemical reaction

## Atoms rearranged

The reactants in a chemical reaction are made up of atoms arranged in a certain way. When a chemical reaction takes place, these atoms are rearranged into a different configuration to form the products. No new atoms are made, and no atoms are destroyed.

Consider the reaction of magnesium metal (Mg) with oxygen gas (O<sub>2</sub>) to form magnesium oxide (MgO). Figure 4.1.3 shows each reactant and product correctly, but there is a big problem.



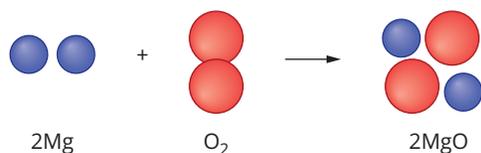
**FIGURE 4.1.3** An equation can be shown by using coloured circles to represent the atoms; this equation is unbalanced as the number of oxygen atoms on the left is different to the number on the right.

In Figure 4.1.3, there are two oxygen atoms on the left side of the arrow, but only one on the right. This equation is therefore unbalanced and must be incorrect, as the oxygen atom cannot simply have disappeared or been destroyed. The extra oxygen atom on the left must be accounted for.

## A balanced approach

In a balanced equation, the numbers of each element on both sides of the equation are the same. This means that no atoms are gained or lost. A balanced equation represents what happens to the atoms in a chemical reaction.

The numbers placed in front of the Mg and the MgO in this equation are called **coefficients**. They indicate the numbers of atoms or molecules of the substances present (Figure 4.1.4).



**FIGURE 4.1.4** In this balanced equation, all the atoms have been accounted for. Atoms are neither made nor destroyed in a chemical reaction; they are simply rearranged. This is a fundamental principle of chemistry, called the **law of conservation of mass**.

## Demonstrating conservation of mass

The law of conservation of mass can be demonstrated by measuring the mass of the reactants, and the mass of the products. Figure 4.1.5 shows silver nitrate solution ( $\text{AgNO}_3$ ) and sodium chloride solution ( $\text{NaCl}$ ) before and after mixing. The products of the reaction—silver chloride solid ( $\text{AgCl}$ ) and sodium nitrate solution ( $\text{NaNO}_3$ )—have the same mass as the reactants. The atoms are unchanged during the reaction process, simply rearranged to form new substances.

### SC 2 CHECK YOUR UNDERSTANDING

State the law of conservation of mass.

**SC 3** I can identify the names or formulas of reactants and products in familiar chemical reactions

There are a lot of different chemical reactions happening around us, and it can sometimes seem difficult to interpret them. Reactions can, however, be classified into different types. Once you are familiar with the various types of reactions, it is much easier to understand the chemistry involved, and even start to make predictions.

## Combustion reactions

In a **combustion** reaction, a fuel reacts with oxygen, producing new products and releasing heat. This release of heat energy means that it is an exothermic reaction.

For more than a century, fossil fuels have been used in combustion reactions to provide energy for society; for example, as coal for heating and cooking, and petrol for driving cars. Fossil fuels are produced from the decay of living things over millions of years.

Two types of combustion reaction are **complete combustion** and **incomplete combustion**. Respiration can also be considered a form of combustion as glucose and oxygen react together to form carbon dioxide and water in human cells.

### KEY TERMS

**coefficient** the big number in front of a chemical in a balanced equation  
**law of conservation of mass** atoms cannot be created or destroyed during a chemical reaction



**FIGURE 4.1.5** The law of conservation of mass: the mass of the reactants in two beakers (left) is the same as the mass of the products (right).

### KEY TERMS

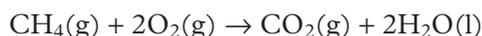
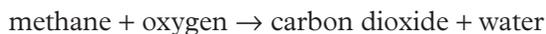
**combustion** a chemical reaction in which a substance burns in oxygen gas to produce light and heat  
**complete combustion** combustion that occurs when there is plenty of oxygen available; it produces carbon dioxide and water vapour  
**incomplete combustion** combustion that occurs when oxygen is limited; produces carbon (soot, smoke) and/or carbon monoxide

## Complete combustion

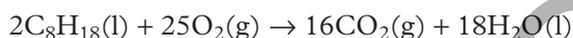
In complete combustion, there is plenty of oxygen to react with the fuel, and the reaction uses as much oxygen as it needs to form carbon dioxide as one of the products. The general formula for complete combustion is:



For example, natural gas, as supplied to homes and industry, is mostly methane. The equations for the complete combustion of methane are:



Another good example is petrol, which is commonly used in vehicles. It is a mixture of many substances, one of which is octane ( $\text{C}_8\text{H}_{18}$ ). The equations for the complete combustion of octane are:



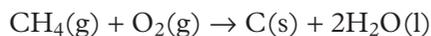
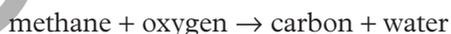
## Incomplete combustion

Sometimes, however, there is a limited supply of oxygen; in this case either carbon monoxide or carbon is produced as one of the products. This is called incomplete combustion. A small deficiency in oxygen will produce carbon monoxide. A greater deficiency in oxygen will produce carbon.

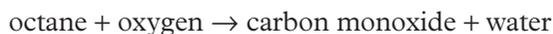
The general formulas for incomplete combustion are:



For example, methane undergoes incomplete combustion in a Bunsen burner when the air hole is closed, and a yellow flame is visible. This limits the amount of oxygen available for the reaction, and produces a dirty flame which releases black carbon (soot) to the surroundings (Figure 4.1.6). The equations for the incomplete combustion of methane in a Bunsen burner are:

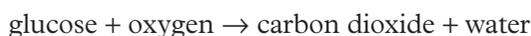


Octane in petrol can also undergo incomplete combustion. This can happen if a car is not serviced correctly and produces carbon monoxide, a serious environmental pollutant:

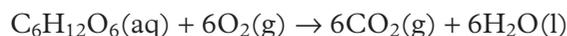


## Respiration

Respiration, which occurs in the cells of living organisms, is a form of combustion, but it does not have a burning flame:



**FIGURE 4.1.6** A Bunsen burner produces a yellow flame and black soot during incomplete combustion.



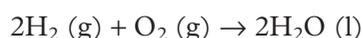
Note the state glucose is in is (aq) or aqueous. This means it is dissolved in water.

## Alternative fuels

Although carbon-based fossil fuels, such as methane and octane, have been widely used in the past, their supply will not last forever. They are classified as **non-renewable** energy sources, as they take millions of years to be produced.

Scientists are, therefore, investigating hydrogen ( $\text{H}_2$ ) to see whether it might be useful in the future as a renewable energy source (Figure 4.1.7). Because hydrogen does not contain carbon, it does not produce carbon dioxide, carbon monoxide or carbon as a product when it undergoes combustion. The equations for the combustion of hydrogen are:

hydrogen + oxygen  $\rightarrow$  water



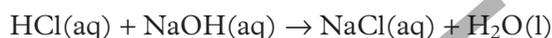
## Neutralisation reactions

When an acid reacts with a **base**, the reaction is called **neutralisation**. The general formula for neutralisation is:

acid + base  $\rightarrow$  salt + water

An example of neutralisation is the reaction of hydrochloric acid (HCl) with sodium hydroxide (NaOH):

hydrochloric acid + sodium hydroxide  $\rightarrow$  sodium chloride + water



Note that the word salt does not always refer to sodium chloride (table salt). Other salts can be formed, depending on the reactants used in the neutralisation reaction. The type of salt produced in a neutralisation reaction will depend on the acid and base used. The metal in the base determines the metal in the salt product (Table 4.1.1). For example, the base magnesium hydroxide, will form a salt with magnesium in it. Some common acids are hydrochloric acid (HCl), sulfuric acid ( $\text{H}_2\text{SO}_4$ ) and nitric acid ( $\text{HNO}_3$ ) (Table 4.1.2).

**TABLE 4.1.1** Name of salts linked to the base reactant produced from a neutralisation reaction

Base (reactant)	Salt has this in its name	Salt has this in its formula	Example of salt product
magnesium hydroxide ( $\text{Mg}(\text{OH})_2$ )	magnesium	Mg	magnesium chloride ( $\text{MgCl}_2$ )
lithium hydroxide (LiOH)	lithium	Li	lithium chloride (LiCl)
potassium hydroxide (KOH)	potassium	K	potassium chloride (KCl)

### KEY TERM

**non-renewable** source of energy that cannot be easily replaced after it is used, such as fossil fuels like coal or gas



**FIGURE 4.1.7** A test model of a hydrogen fuelled car

### KEY TERMS

**base** a substance that will neutralise an acid

**neutralisation** a reaction of an acid with a base, forming a salt and water

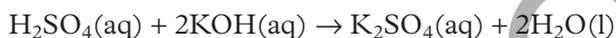
**TABLE 4.1.2** Name of salts linked to the acid reactant produced from a neutralisation reaction.

Acid (reactant)	Type of salt formed	Salt has this in its formula	Example of salt
Hydrochloric acid (HCl)	chloride	Cl	sodium chloride (NaCl)
sulfuric acid (H <sub>2</sub> SO <sub>4</sub> )	sulfate	SO <sub>4</sub>	sodium sulfate (Na <sub>2</sub> SO <sub>4</sub> )
nitric acid (HNO <sub>3</sub> )	nitrate	NO <sub>3</sub>	potassium nitrate (KNO <sub>3</sub> )

When naming the salt, use the metal from the base first, followed by the word associated with the acid.

For example, the reaction of sulfuric acid with potassium hydroxide is:

sulfuric acid + potassium hydroxide → potassium sulfate + water



## Metal-acid reactions

The general formula for the reaction between a metal and an acid is:

acid + metal → salt + hydrogen

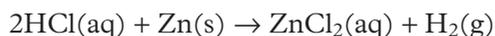
The salts produced can be worked out similar to the ones used for neutralisation reactions by referring to Table 4.1.3.

**TABLE 4.1.3** Examples of a metal-acid reaction

Acid	Type of salt	Salt has this in its formula	Example of salt
hydrochloric acid (HCl)	chloride	Cl	calcium chloride (CaCl <sub>2</sub> )
sulfuric acid (H <sub>2</sub> SO <sub>4</sub> )	sulfate	SO <sub>4</sub>	potassium sulfate (K <sub>2</sub> SO <sub>4</sub> )
nitric acid (HNO <sub>3</sub> )	nitrate	NO <sub>3</sub>	magnesium nitrate (Mg(NO <sub>3</sub> ) <sub>2</sub> )

An example would be the reaction between hydrochloric acid and zinc metal:

hydrochloric acid + zinc → zinc chloride + hydrogen



Note that nitric acid does not always form a salt and hydrogen when it reacts with metals.

Only the more reactive metals will react with acids in this way. For example, unreactive metals such as copper, silver and gold will not react with acids to produce a salt and hydrogen.

**SC 3 CHECK YOUR UNDERSTANDING**

Identify the products in the following chemical reactions by writing out the chemical formulas:

- a  $2\text{NaOH}(\text{aq}) + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{Na}_2\text{SO}_4(\text{aq}) + 2\text{H}_2\text{O}(\text{l})$
- b  $2\text{Fe}(\text{s}) + 6\text{HCl}(\text{aq}) \rightarrow 2\text{FeCl}_3(\text{aq}) + 3\text{H}_2(\text{g})$
- c  $2\text{C}_4\text{H}_{10}(\text{aq}) + 13\text{O}_2(\text{g}) \rightarrow 8\text{CO}_2(\text{g}) + 10\text{H}_2\text{O}(\text{l})$

## Lesson review

Use these questions to check whether you have met the learning intention for this lesson.

- 1 Write the chemical formulas of the reactants and products in the following reaction:  
 $\text{NaCl}(\text{aq}) + \text{AgNO}_3(\text{aq}) \rightarrow \text{NaNO}_3(\text{aq}) + \text{AgCl}(\text{s})$
- 2 Explain how the amount of oxygen available for combustion results in complete or incomplete combustion.
- 3 A student conducts an experiment where they react hydrochloric acid with a metal. The total mass of the reactants was 5.045 g. The student adds the metal to an open beaker containing the acid. They observe bubbles and fizzing during the reaction. After the reaction, they weigh the products in the beaker and the mass was 3.089 g. The student concludes that the conservation of mass did not occur.  
Explain what is wrong with their observation.
- 4 Consider the reaction  $2\text{Al}(\text{s}) + 3\text{Cl}_2(\text{g}) \rightarrow 2\text{AlCl}_3(\text{s})$ .
  - a Calculate the total mass of reactants if you start with 54 g of Al and 213 g of  $\text{Cl}_2$ .
  - b Determine the total mass of products.
  - c Explain how this calculation supports the law of conservation of mass.

## 4.2

# Conservation of mass in a chemical reaction

## Learning intention

To be able to measure and record accurate data to provide evidence for the conservation of mass in a chemical reaction

## Success criteria

**SC 1:** I can measure and record data from an investigation accurately.

**SC 2:** I can minimise sources of error when collecting data.

**SC 3:** I can use data from an investigation to support a prediction related to the theory of conservation of mass.



**FIGURE 4.2.1** An example of a precipitation reaction, where two clear solutions mix together and a yellow solid forms.

### KEY TERMS

**precipitation reaction** when two clear solutions react to produce an insoluble solid

**precipitate** a solid formed during a chemical reaction

## Introduction

One of the most important laws in science is that matter cannot be created or destroyed in a chemical reaction. According to this law, all the atoms present in the reactants of a chemical reaction will still be present in the products of that reaction.

In this practical investigation, you will test the law of conservation of mass by measuring the mass of chemicals before and after a range of precipitation reactions (Figure 4.2.1).

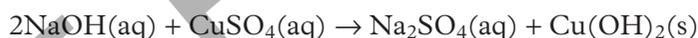
## Background

Chemical reactions can be classified under a range of types, including combustion, neutralisation and precipitation. In **precipitation reactions**, two solutions are added together and a solid is formed, normally instantly. The solid formed is not soluble in water and is called the **precipitate**. The colour of the precipitate will depend on the chemicals present in the original solutions.

An example of a precipitation reaction that you will use in this investigation is sodium hydroxide solution being added to a solution of copper sulfate.

The word and symbol equations for the reaction are shown here:

sodium hydroxide + copper sulfate → sodium sulfate + copper hydroxide



Note that the (aq) state symbol means that this is an aqueous solution in water, and the (s) means the chemical is a solid.

## Aim

To demonstrate that mass is conserved in a chemical reaction

## Prediction

Read the method for the investigation and write a prediction for the three experiments based on the law of conservation of mass.

## Materials

- 60 mL of 0.1 M copper sulfate solution
- 20 mL of 0.1 M sodium hydroxide solution
- 20 mL of 0.1 M sodium carbonate solution
- 20 mL of 0.1 M ammonia solution

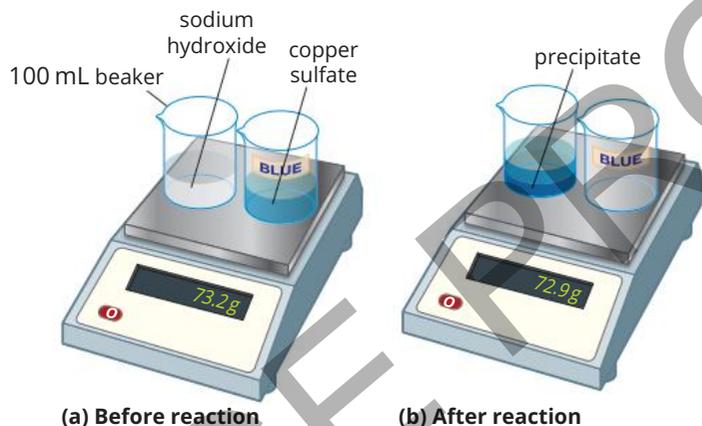
- four 100 mL beakers
- marker pen
- electronic balance

## Assessment of risk

Ensure you are aware of the risks of this practical investigation and have considered how safety can be improved before carrying out this activity.

## Method

- 1 Create a results table to record the mass of the beakers and solutions before and after the reaction for each of the three experiments in this investigation.
- 2 Pour copper sulfate solution into one of the beakers until it reaches the 20 mL mark. Use the marking pen to label this beaker BLUE.
- 3 To the other beaker, add 20 mL of sodium hydroxide solution.
- 4 Place both beakers on the electronic balance and determine their total mass as shown in Figure 4.2.2. Record this mass.



**FIGURE 4.2.2** A simplified diagram demonstrating how to weigh the beakers.

- 5 Carefully pour the copper sulfate solution into the sodium hydroxide solution and observe what happens.
- 6 Place the beaker with the solutions in it and the empty beaker labelled BLUE back on the scales as shown in the after reaction diagram. Record the total mass.
- 7 For the second experiment, add 20 mL of the sodium carbonate solution to a clean beaker. Add another 20 mL of copper sulfate solution to the beaker labelled BLUE. Record the total mass.
- 8 Carefully pour the copper sulfate solution into the sodium carbonate solution. As before, record the total mass of the beakers after the reaction.
- 9 For the third experiment, repeat steps 7 and 8, but this time use the ammonia solution instead of the sodium carbonate solution.

## SAFETY NOTES

- ▶ All chemicals should be treated as toxic.
- ▶ Do not dispose of any of the solutions by pouring them down the sink. Follow your teacher's instructions for the disposal of solutions used in this experiment.
- ▶ Do not taste any chemicals.

## Results

Record your results in a table.

Reactants	Total mass before mixing (g)	Total mass after mixing (g)	Observations
sodium hydroxide + copper sulfate			
sodium carbonate + copper sulfate			
ammonia + copper sulfate			

## Conclusion

- 1 What evidence shows that a chemical reaction took place once the solutions were mixed?
- 2 Compare the total masses before and after mixing.
- 3 State the law of conservation of mass and assess whether your results support this law.

### GO TO

SkillBuilder: Reducing error in experiments in your Skills Toolkit, page xx and Worked example: Sources of error, page xx

## Evaluation

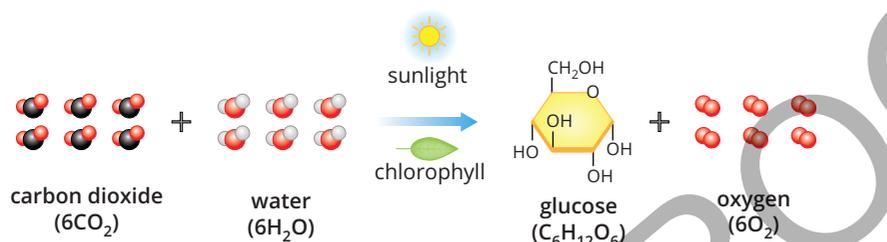
- 1 List two things that you did to ensure the results were accurate.
- 2 Name one source of random error and one source of systematic error that could affect the results in this investigation.

## 4.3 Types of chemical reactions

### Lesson overview

Oxygen is a powerful element. Oxygen gas ( $O_2$ ) is a colourless, odourless gas and makes up 21% of Earth's atmosphere by volume, while oxygen atoms comprise 49% of the mass of Earth's crust. Oxygen is produced naturally by photosynthesis in plants (Figure 4.3.1), and all living things use oxygen to release energy for their survival in respiration. Furthermore, oxygen is widely used in manufacturing, welding and in the treatment of sewage.

In this lesson, you will learn about combustion, and how oxygen enables things to burn. You will also learn about corrosion. Many corrosion reactions involve oxygen, but you will also explore the role of other elements in the corrosion of various metals.



**FIGURE 4.3.1** Photosynthesis uses carbon dioxide and water as reactants to form glucose and oxygen as products.

**SC 1** I can describe combustion and corrosion reactions with word equations

### Combustion reactions

When a substance reacts with oxygen gas to produce heat and light, it is undergoing combustion. All combustion reactions are exothermic, but they don't all occur at the same rate.

Fast combustion is violently explosive. For example, petrol in a car engine reacts rapidly to drive the piston. Slow combustion is where the reaction is slower, more controlled. Gas stoves are an example of slow combustion.

When writing word equations for combustion reactions, make sure you know if the reaction is undergoing complete combustion or incomplete combustion.

Complete combustion has carbon dioxide and water as their products.

carbon-based fuel + oxygen  $\rightarrow$  carbon dioxide + water

Incomplete combustion has carbon monoxide or carbon as well as water as their products.

carbon-based fuel + oxygen  $\rightarrow$  carbon monoxide + water

carbon-based fuel + oxygen  $\rightarrow$  carbon + water

### Corrosion reactions

Sometimes a material will react with other chemicals in its immediate environment. These chemicals consume some of the material's surface and form a new substance. This process is called **corrosion**.

#### Learning intention

To understand that similarities exist between different types of chemical reactions

#### Success criteria

**SC 1:** I can describe combustion and corrosion reactions with word equations.

**SC 2:** I can describe differences and similarities between corrosion and combustion reactions.

**SC 3:** I can describe how environmental conditions can affect combustion and corrosion reactions.

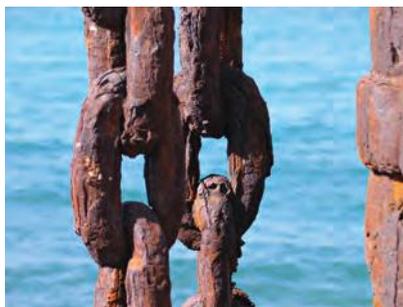
#### Scifile

##### Colourful combustion

Fireworks are a spectacular example of combustion reactions. Different chemicals produce different colours when they burn, making the night sky light up with a variety of hues.

#### KEY TERM

**corrosion** when a metal reacts with the chemicals in its environment



**FIGURE 4.3.2** The iron in this shipwreck has corroded (rusted) extensively.

### KEY TERMS

**rust** hydrated iron(III) oxide; chemical formula  $\text{Fe}_2\text{O}_3 \cdot \text{H}_2\text{O}$

**hydrated** a substance that contains water

**verdigris** a bluish-green layer that forms on copper, brass or bronze due to reaction with oxygen, water and carbon dioxide

**tarnish** a black coating of silver sulfide that is produced when silver reacts with sulfur in food or the atmosphere; chemical formula  $\text{Ag}_2\text{S}$



**FIGURE 4.3.4** The Statue of Liberty, covered in its verdigris of copper(II) hydroxide and copper(II) carbonate.

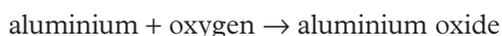


**FIGURE 4.3.5** The candlestick on the right shows the lustrous silver metal, while the one on the left has tarnished.

## Corrosion by oxygen

Some metals can react with oxygen without releasing light energy and, despite being exothermic, do not cause a rapid increase in temperature. When this happens, an oxide (a compound of metal and oxygen) of the metal is formed. This is an example of corrosion.

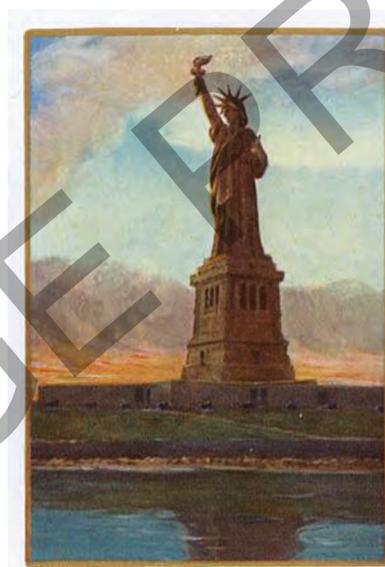
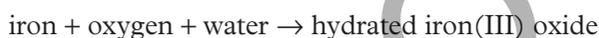
Often the oxide formed is durable and stable and does not cause damage to the metal. For example, aluminium reacts with oxygen to form a layer of hard aluminium oxide, which protects the aluminium underneath. This is why objects made from aluminium, such as window frames, can last a long time. The word equation for this reaction is:



## Corrosion by oxygen and water

Iron reacts with oxygen and water to form a flaky red-brown oxide called **rust**. As the rust flakes off, the iron underneath will continue to react, leading to more corrosion, or disintegration. An unprotected piece of iron can corrode completely over time (Figure 4.3.2).

The word equation for the rusting of iron is:



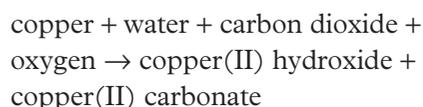
**FIGURE 4.3.3** A coloured playing card of the Statue of Liberty from early 1900s which shows the copper before it started to corrode.

The word **hydrated** means that a substance contains water (it has absorbed water). Just as iron oxide is a compound of iron and oxygen, hydrated iron(III) oxide (rust) is a compound of iron, oxygen, and water.

## Corrosion by oxygen, water and carbon dioxide

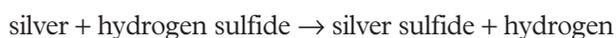
Copper can react with oxygen, water, and carbon dioxide to produce a green substance called verdigris. A good example of this is the Statue of Liberty in New York City. It was a red-brown colour originally (Figure 4.3.3), but has slowly changed to the blue-green **verdigris** – a combination of copper(II) hydroxide and copper(II) carbonate – over time (Figure 4.3.4).

The word equation for this reaction is:



## Corrosion by hydrogen sulfide

Silver can react with hydrogen sulfide, which is present in small amounts in the air, to produce **tarnish**. Silverware, when tarnished, gets coated in a layer of black silver sulfide (Figure 4.3.5). The word equation for this reaction is:





## SCIENCE IN SOCIETY

### Protecting metal from corrosion

Steel is made from iron melted together with small amounts of carbon or other metals. When a metal is mixed with either carbon or another metal, it is known as an alloy. Alloying iron with carbon, to make steel, increases the strength of the iron and can change other physical properties.

Steel has countless uses. For example, it is made into common household products, such as cutlery and saucepans, medical instruments, such as scalpels, industrial and commercial products, such as cutting tools, wires, screws, nails and vehicle components. Larger scale uses for steel include train tracks, buildings, ships and trucks as well as infrastructure like bridges. For example, the iconic Sydney Harbour Bridge is made of steel (Figure 4.3.6).



**FIGURE 4.3.6** Sydney Harbour Bridge is made from steel; it was built in 1920s and is regularly inspected for corrosion.

With all these vital products and infrastructure containing steel, how do you prevent or slow down the corrosion? Corrosion will reduce the quality of the metal, its strength, visual aesthetics, durability and other physical properties.

Some metals, such as chromium and nickel are resistant to corrosion and so are commonly used in steel alloys to slow down the corrosion. Chromium and nickel react with oxygen to form a protective chromium oxide layer on the surface. Nickel and chromium are mixed with steel to form stainless steel – the product you will find in most kitchen items like cutlery, sinks and saucepans. And commercial kitchens are made from stainless steel (Figure 4.3.7). The portions of nickel and chromium added to the steel is depended on the use of the stainless steel (as there are over 100 different varieties!).



**FIGURE 4.3.7** Commercial kitchens are made from stainless steel.

There are other ways steel can be protected from corrosion. Some options include:

- **Galvanising.** Galvanised steel is where steel is dipped into molten zinc, so a protective coat of zinc covers the steel. Zinc corrodes with oxygen and carbon dioxide to form a protective surface layer. Trailers are commonly galvanised.
- **Coating with a corrosion resistant metal.** Tin cans, in the past, were steel cans coated with a thin layer of tin. Tin protects the steel from corrosion. Nowadays, most tin cans are coated in plastic instead of tin.
- **Paint.** Painting is a very common method of protecting steel from corrosion, from large objects like the Sydney Harbour Bridge to cars.
- **Coating with a polymer.** Some non-stick frying pans are steel, coated with a non-stick polymer that protects the steel as well as makes it easy to cook and clean.
- **Powder coatings.** Powder coatings are fine particles of pigment and resin which is sprayed onto the surface of the steel and heated to fuse the particles to the metal. Powder coated steel are typically used for garage doors, fences, sheds and roofs (Figure 4.3.8).



**FIGURE 4.3.8** Powder-coated steel used for fences.

### SC 1 CHECK YOUR UNDERSTANDING

Name the type of reaction represented by this word equation:



**SC 2** I can describe differences and similarities between corrosion and combustion reactions

Corrosion reactions are similar, in some ways, to combustion reactions. But they are not the same and the important differences are summarised in Table 4.3.1.

**TABLE 4.3.1** Comparison of corrosion and combustion reactions

Combustion	Corrosion
	
Rapid chemical change	Slow chemical change
Heat and light energy are released	Heat energy is released gradually, with no measurable temperature change
Sometimes the energy release is explosive	The products form a new surface on the affected material
The affected material is known as the fuel	This surface can be either hard and protective or flaky and damaging
Oxygen is always one of the reactants	Various substances can cause corrosion, including oxygen, water, carbon dioxide, and hydrogen sulfide
<b>Similarities:</b> Both are exothermic, consume the affected material, and cause a colour change	

### SC 2 CHECK YOUR UNDERSTANDING

List the reactants that cause corrosion and combustion of a metal.

**SC 3** I can describe how environmental conditions can affect combustion and corrosion reactions

## Combustion reactions

During a combustion reaction, the environmental conditions determine whether it will be complete or incomplete.

When many fuels are burnt in the presence of abundance of oxygen, complete combustion occurs to produce carbon dioxide and water. This is called efficient burning. Appliances such as gas cookers are designed to burn efficiently in order to avoid dirtying cookware with soot (carbon) or introducing poisonous gases (carbon monoxide).

However, when oxygen is restricted, fuels can burn inefficiently. For instance, the cylinder of an internal combustion engine in an automobile has a limited oxygen supply and produces some carbon monoxide. For this reason, adequate ventilation is important in car parks and workshops.

Sometimes the fuel itself can affect its combustion products. Tyres, for example, contain a mixture of different molecules, including some very large ones known as **polymers**. One of the polymers in tyres is polybutadiene, which burns to produce a dirty black smoke consisting of carbon (Figure 4.3.9).

polybutadiene + oxygen → carbon + water

## Corrosion reactions

Environmental conditions are also critical for corrosion reactions, as shown in the following examples.

### Aluminium

Some forms of corrosion, such as the formation of the hard layer of aluminium oxide on the surface of aluminium metal, happen readily in the presence of oxygen. If this protective layer is chemically removed, the exposed aluminium reacts readily in moist air, causing further damage to the metal (Figure 4.3.10).

The high reactivity of aluminium explains why it is not found in Earth as a pure metal but always combined with other elements in the form of aluminium compounds. Therefore, aluminium must be extracted from aluminium ore, a rock containing an aluminium compound. There will be more about the reactivity of metals and their extraction from Earth in a later lesson.

### Iron

Rusting of iron requires the presence of oxygen and water. Without one or the other, the reaction cannot proceed. Water usually has some oxygen dissolved in it.

Rusting of iron also requires the presence of an **electrolyte**, which is a substance that can dissolve in water to produce a solution that conducts electricity. This is important because the rusting reaction involves small electrical currents passing through the water. Most water contains the electrolyte carbonic acid, which is formed when carbon dioxide dissolves in water. When a solution of carbonic acid is present, the products of the corrosion can include carbonates and hydroxides. A much stronger electrolyte is sodium chloride (NaCl). This is why iron rusts to a much greater extent close to the ocean, where sodium chloride (salt) from the sea dissolves in water in the environment.

Boiling water removes the oxygen and carbon dioxide dissolved in it. If a clean iron nail is placed in boiled water (with no salts) under a layer of oil, which prevents gases from dissolving back into the water, it will not rust. Although water is present, there is no oxygen and no electrolyte. If the same iron nail is placed into a **desiccant**, which removes the water from the atmosphere, it will not rust either. Although oxygen is present, there is no water. See Figure 4.3.11 as a demonstration of iron corrosion.

#### SC 3 CHECK YOUR UNDERSTANDING

Explain why iron in boiled water does not rust as easily as iron in tap water.



**FIGURE 4.3.9** Friction from drifting during this racing event causes the tyres to burn, releasing smoke made from solid carbon.



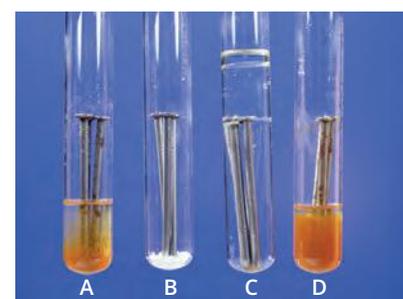
**FIGURE 4.3.10** The piece of aluminium on the left has formed a layer of aluminium oxide, which protects it; the one on the right has had this layer removed, causing it to react with moist air.

#### KEY TERM

**polymer** a very large molecule composed of many repeating subunits

**electrolyte** a substance that can dissolve to form a solution that conducts electricity and/or can conduct electricity in its molten form

**desiccant** a substance that absorbs water from the surrounding air



**FIGURE 4.3.11** The iron in tube A is exposed to air and water so it rusts; the iron in tube B is kept dry with a desiccant so it does not rust; tube C contains boiled water and is sealed from the atmosphere with a layer of oil, so the iron does not rust; tube D contains salt water, which has increased the extent of rusting.

## Lesson review

Use these questions to check whether you have met the learning intention for this lesson.

- Classify the following word equations as combustion or corrosion:
  - zinc + oxygen  $\rightarrow$  zinc oxide
  - ethane + oxygen  $\rightarrow$  carbon dioxide + water
  - sodium + oxygen  $\rightarrow$  sodium oxide
  - wood + oxygen  $\rightarrow$  carbon + water
- Consider the following reactions:
  - Write the word equation for the complete combustion of methane ( $\text{CH}_4$ ).
  - Write the word equation for the corrosion of magnesium.
  - Compare the reactants in both equations and describe the similarities.
  - Compare the products in both equations and describe the differences.
- When silver tarnishes, a black compound forms on the surface of the silver.
  - Name the substance silver reacts with when it tarnishes.
  - Name the silver compound that is made in this reaction.
- Countries that have snow every winter spread large quantities of salt (typically  $\text{NaCl}$  or table salt) on their roads to reduce the risk of car accidents due to ice. However, a disadvantage of this safety measure is that cars rust more quickly in those countries. Using your knowledge of corrosion, explain this phenomenon.



- Classify the following statements according to their application for corrosion only or combustion only.
  - heat and light energy is released
  - oxygen is always one of the reactants
  - the product forms a new surface on the metal
  - heat is released gradually with no measurable change in temperature
  - carbon dioxide is a reactant

## 4.4 Comparing chemical reactions

### Introduction

Hydrochloric acid is a substance commonly used in the production of cleaners and plastics, in food processing, and in many other industrial applications. It is also present inside the stomach where it helps to digest food. This acid is so useful because it can react with a range of substances, and these reactions can be compared to identify similarities and differences from which patterns emerge.

In this practical investigation, you will study four important reactions of hydrochloric acid, for example, adding a metal carbonate (Figure 4.4.1). Applying your knowledge from earlier lessons, plus some simple chemical tests, you will identify the products of these reactions and look for trends in your results.

### Background

Hydrochloric acid can react to produce various products depending on the other reactant. Although it can seem difficult to remember what the products of these chemical reactions will be, this task is made a lot easier if you remember that reactions form patterns. These patterns are known as general equations.

The general equations needed for this investigation are:

acid + metal oxide  $\rightarrow$  salt + water

acid + reactive metal  $\rightarrow$  hydrogen + salt

acid + metal carbonate  $\rightarrow$  salt + carbon dioxide + water

acid + metal hydrogen carbonate  $\rightarrow$  salt + carbon dioxide + water

As part of this investigation, you will be testing the products of each reaction to see whether the expected gases were produced. For example, limewater is used to test for the presence of carbon dioxide and a lit match is used to test for the presence of hydrogen which will ignite to produce an audible 'pop'. Limewater is a saturated solution of calcium hydroxide. Limewater reacts with carbon dioxide to form a white precipitate of calcium carbonate.

### Aim

To investigate some common reactions of hydrochloric acid

### Hypothesis

Hydrochloric acid solution will react with metal oxides, reactive metals, metal carbonates, and metal hydrogen carbonates.

#### Learning intention

To be able to carry out, describe, represent and compare a range of chemical reactions

#### Success criteria

**SC 1:** I can identify and describe evidence for chemical change in a range of reactions.

**SC 2:** I can write word equations for observed chemical reactions.

**SC 3:** I can compare types of observed chemical reactions.



**FIGURE 4.4.1** What might happen when this metal carbonate is added to hydrochloric acid solution?

**SAFETY NOTES**

- ▶ Wear laboratory coats and safety glasses at all times.
- ▶ Avoid contact with the acid.
- ▶ Use the magnesium ribbon only as directed in the method and keep away from any flame.
- ▶ Follow the directions of your teacher or laboratory manager for the disposal of copper(II) oxide.

## Materials

- 1 M hydrochloric acid (HCl) solution
- calcium carbonate (CaCO<sub>3</sub>) powder
- copper(II) oxide (CuO) powder
- sodium hydrogen carbonate (NaHCO<sub>3</sub>) powder
- 1 cm strip of clean magnesium (Mg) ribbon
- limewater
- 5 test tubes
- dropping pipette
- test-tube rack
- rubber stopper
- long matches or wooden splints
- one-hole stopper and delivery tube or sidearm test tube with rubber stopper and delivery tube

## Assessment of risk

Ensure you are aware of the risks of this practical investigation and have considered how safety can be improved before carrying out this activity.

## Method

- 1 Add 5 mL of 1 M hydrochloric acid (HCl) solution to a small amount of copper(II) oxide (CuO) powder in a test tube. Carefully shake the test tube from side to side to mix the reactants.
- 2 Allow the test tube to stand for 5 minutes before recording your observations in the results table.
- 3 Add 5 mL of 1 M HCl solution to a second test tube containing a 1 cm strip of magnesium ribbon. Place a stopper lightly on top of the test tube and collect gas for 10 seconds. (Do not put the stopper on tightly, as the test tube might crack.)
- 4 Hold the test tube at an angle of around 45° with the neck facing away from you.
- 5 When the fizzing stops, remove the stopper and test the gas with a lit match or burning wooden splint. Record your observations in the results table.
- 6 Fill a third test tube to a depth of about one third with limewater.
- 7 Add 10 mL of 1 M HCl solution to a test tube containing a small quantity of calcium carbonate (CaCO<sub>3</sub>) powder and quickly add a stopper with a delivery tube connected.
- 8 Place the end of the delivery tube into the limewater. Record your observations in the results table.
- 9 Repeat steps 7 and 8 but replace the calcium carbonate with sodium hydrogen carbonate (NaHCO<sub>3</sub>) powder. Record your observation in the results table.

## Results

Record your experimental results in a table like this.

Test number	Reactants	Test performed	Observations
1	HCl + CuO	lit match	
2	HCl + Mg	lit match	
3	HCl + CaCO <sub>3</sub>	limewater	
4	HCl + NaHCO <sub>3</sub>	limewater	

## Conclusion

- 1 In a table like the one below, state whether a chemical change has occurred, and briefly describe the experimental evidence.

Test number	Reactants	Chemical change?	Evidence for chemical change
1	HCl + CuO		
2	HCl + Mg		
3	HCl + CaCO <sub>3</sub>		
4	HCl + NaHCO <sub>3</sub>		

- 2 State which product can be identified using limewater.
- 3 State which product can be identified using a lit match (the pop test).
- 4 Based on your experimental results and using your answers to questions 2 and 3 above, write a word equation for each of the reactions investigated.
- 5 Four different reactions with an acid were considered in this investigation. Based on the information in the Background, copy and complete the table below to summarise their similarities and differences.

Reaction	Salt produced? (yes or no)	Water produced? (yes or no)	Gas produced (identify gas)
acid + metal oxide (HCl + CuO)			
acid + reactive metal (HCl + Mg)			
acid + metal carbonate (HCl + CaCO <sub>3</sub> )			
acid + metal hydrogen carbonate (HCl + NaHCO <sub>3</sub> )			

- 6 Write an overall conclusion for this investigation.

## Evaluation

This practical investigation established that hydrochloric acid reacts with four different compounds. Each reaction produced a distinctive set of products, though the investigation only involved hydrochloric acid and one example of a metal oxide, reactive metal, metal carbonate, and metal hydrogen carbonate.

Complete an evaluation of this experiment and suggest how it could be modified to allow for a more thorough investigation of whether the general formulas for acid reactions hold true.

## 4.5

# Atomic rearrangement in chemical reactions

## Learning intention

To understand that atomic rearrangement occurs when reactants are converted into products during a chemical reaction

## Success criteria

**SC 1:** I can identify specific atoms found in the reactants and products of chemical equations.

**SC 2:** I can use chemical formulas to calculate the quantity of specific atoms found in the reactants and products of chemical reactions.

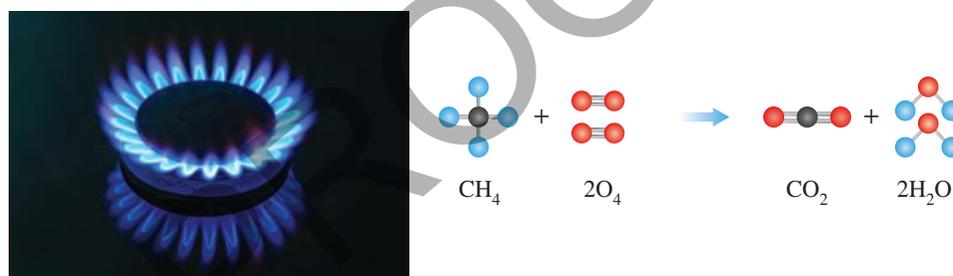
**SC 3:** I can predict the products formed in a chemical reaction based on the atomic arrangement in the reactants.

## Lesson overview

In chemical reactions, the atoms in reactants are rearranged to form new chemicals called products.

Chemical equations provide a convenient way of displaying what happens in a chemical reaction, as they represent the substances that react and form using chemical formulas. A chemical formula comprises the symbols of the elements in the substance as well as numbers to show how many atoms of each element are present (Figure 4.5.1).

In this lesson, you will learn to identify the atoms shown in chemical equations and understand how chemical formulas describe the number of atoms of each element in the reactants and products of chemical reactions. You will then apply this knowledge to predict the products of a chemical reaction.



**FIGURE 4.5.1** The burning of methane involves the rearrangement of atoms to form the compounds carbon dioxide and water.

**SC 1** I can identify specific atoms found in the reactants and products of chemical equations

## Identifying elements in chemical equations

All known atoms and their symbols are listed on the periodic table of elements (Figure 4.5.2).

Each of the 118 elements has a unique name, chemical symbol and atomic number. Typically, chemical symbols are made up of either a single capital letter, such as fluorine (F), or a single capital letter followed by a lowercase letter, such as aluminium (Al).

Most chemical symbols are relatively obvious; for example, the symbol for hydrogen is H and the symbol for chlorine is Cl. However, the chemical symbols of some elements are less clear as they are derived from another language, usually Latin. For example, the chemical symbol for lead is Pb, derived from *plumbum*, and the chemical symbol for iron is Fe, from *ferrum*.

KEY																						
1 H hydrogen																2 He helium						
3 Li lithium		4 Be beryllium														5 B boron		6 C carbon	7 N nitrogen	8 O oxygen	9 F fluorine	10 Ne neon
11 Na sodium		12 Mg magnesium														13 Al aluminium	14 Si silicon	15 P phosphorus	16 S sulfur	17 Cl chlorine	18 Ar argon	
19 K potassium	20 Ca calcium	21 Sc scandium	22 Ti titanium	23 V vanadium	24 Cr chromium	25 Mn manganese	26 Fe iron	27 Co cobalt	28 Ni nickel	29 Cu copper	30 Zn zinc	31 Ga gallium	32 Ge germanium	33 As arsenic	34 Se selenium	35 Br bromine	36 Kr krypton					
37 Rb rubidium	38 Sr strontium	39 Y yttrium	40 Zr zirconium	41 Nb niobium	42 Mo molybdenum	43 Tc technetium	44 Ru ruthenium	45 Rh rhodium	46 Pd palladium	47 Ag silver	48 Cd cadmium	49 In indium	50 Sn tin	51 Sb antimony	52 Te tellurium	53 I iodine	54 Xe xenon					
55 Cs caesium	56 Ba barium	57-71 lanthanoids		72 Hf hafnium	73 Ta tantalum	74 W tungsten	75 Re rhenium	76 Os osmium	77 Ir iridium	78 Pt platinum	79 Au gold	80 Hg mercury	81 Tl thallium	82 Pb lead	83 Bi bismuth	84 Po polonium	85 At astatine	86 Rn radon				
87 Fr francium	88 Ra radium	89-103 actinoids		104 Rf rutherfordium	105 Db dubnium	106 Sg seaborgium	107 Bh bohrium	108 Hs hassium	109 Mt meitnerium	110 Ds darmstadtium	111 Rg roentgenium	112 Cn copernicium	113 Uut ununtrium	114 Fl flerovium	115 Uup ununpentium	116 Lv livermorium	117 Uus ununseptium	118 Uuo ununoctium				
Lanthanoids		57 La lanthanum	58 Ce cerium	59 Pr praseodymium	60 Nd neodymium	61 Pm promethium	62 Sm samarium	63 Eu europium	64 Gd gadolinium	65 Tb terbium	66 Dy dysprosium	67 Ho holmium	68 Er erbium	69 Tm thulium	70 Yb ytterbium	71 Lu lutetium						
Actinoids		89 Ac actinium	90 Th thorium	91 Pa protactinium	92 U uranium	93 Np neptunium	94 Pu plutonium	95 Am americium	96 Cm curium	97 Bk berkelium	98 Cf californium	99 Es einsteinium	100 Fm fermium	101 Md mendelevium	102 No nobelium	103 Lr lawrencium						

FIGURE 4.5.2 The periodic table displaying all known and named elements as of 2024.

## Elements in reactions

Chemical symbols are used to write chemical equations, so it is important to be able to identify the elements (and their corresponding symbols) in these equations correctly. For example, the incomplete combustion of methane is described in the following equation:

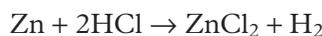


Note that the reaction produces the compound carbon monoxide, CO, and not the element cobalt, Co. Therefore, correct capitalisation is very important.

The reactions that you will explore in this unit include acid reactions with metals, bases and carbonates (or hydrogen carbonates). Two examples are shown below.

### Example 1: Reaction of a metal with an acid

Zinc (Zn) reacts with hydrochloric acid (HCl) to produce zinc chloride (ZnCl<sub>2</sub>) and hydrogen gas H<sub>2</sub>.



Hydrochloric acid (HCl) contains two capital letters, which means it contains two elements: hydrogen (H) and chlorine (Cl). It must not be written as HCL.

The product zinc chloride (ZnCl<sub>2</sub>) also has two capital letters, showing that it contains the two elements zinc (Zn) and chlorine (Cl). The other product just contains hydrogen (H), with two hydrogen atoms in every molecule of hydrogen gas.

## Example 2: Reaction of a hydrogen carbonate with an acid

Phosphoric acid,  $\text{H}_3\text{PO}_4$  reacts with nickel hydrogen carbonate,  $\text{Ni}(\text{HCO}_3)_2$  and produces nickel(II) phosphate ( $\text{Ni}_3(\text{PO}_4)_2$ ), water ( $\text{H}_2\text{O}$ ) and carbon dioxide ( $\text{CO}_2$ ).

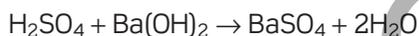


Phosphoric acid ( $\text{H}_3\text{PO}_4$ ) has three capital letters: H, P and O. Since there are no lowercase letters present, these capitals represent the three elements hydrogen (H), phosphorus (P) and oxygen (O) respectively. PO should not, therefore, be confused with the element polonium (Po), which has a lowercase o.

The other reactant, nickel hydrogen carbonate ( $\text{Ni}(\text{HCO}_3)_2$ ), has five letters, four of which are capitals. Therefore, the four elements present are nickel (Ni), hydrogen (H), carbon (C) and oxygen (O). Again, Ni represents nickel and not nitrogen and iodine, which would be represented by the letters NI.

### SC 1 CHECK YOUR UNDERSTANDING

Identify the elements in the following reaction.



### SC 2

I can use chemical formulas to calculate the quantity of specific atoms found in the reactants and products of chemical reactions

## Chemical equations

Chemical equations, such as those presented earlier in this topic, are written with the substances that react together (reactants) written to the left of the arrow and the substances produced (products) written to the right of the arrow.

reactants  $\rightarrow$  products

To illustrate the parts of a balanced chemical equation, consider the formation of sulfur trioxide from sulfur dioxide and oxygen shown below.



You will learn more about balancing chemical equations later in the topic, so for now just focus on the individual substances within the equation.

## Chemical formulas

**Chemical formulas** are a shorthand representation of a substance using chemical symbols, as shown below. They are made up of three basic parts—the coefficient (green), the symbol (blue) and subscripts (red)—each of which have a very specific meaning (Figure 4.5.3).

### KEY TERM

**chemical formula** a representation that uses symbols to indicate the elements in a substance and the relative



**FIGURE 4.5.3** Chemical formula of sulfur dioxide from the equation above showing the coefficient (green), symbols (blue) and subscript (red).

## Reading chemical equations

Table 4.5.1 shows the parts that make up a chemical equation and how to read them.

**TABLE 4.5.1** How to read chemical equations

Definition	What they represent in a chemical equation	Using this reaction as example: $2\text{SO}_2 + \text{O}_2 \rightarrow 2\text{SO}_3$
<b>Coefficients</b>		
numbers in front of the substances	how many molecules of the substance are involved in the chemical equation	$\text{SO}_2$ and $\text{SO}_3$ each have the coefficient <b>2</b> . There is no coefficient number in front of the $\text{O}_2$ . This means the molecule has the coefficient of 1. The number 1 is not used.
<b>Symbols</b>		
each element starts with a capital letter	the elements present in the formula	Sulfur dioxide and sulfur trioxide each contain sulfur ( <b>S</b> ) and oxygen ( <b>O</b> ) atoms. An oxygen molecule contains just oxygen atoms.
<b>Subscripts</b>		
the smaller numbers written after the atoms they refer to	the relative numbers of atoms present	$\text{SO}_2$ : the subscripted <b>2</b> after the oxygen atom means that sulfur dioxide contains one sulfur atom and two oxygen atoms. The number 1 is not used.

## Polyatomic ions

You will also see **polyatomic ions** in some chemical equations. These are compounds, containing a group of atoms, which have a positive or negative electric charge (These charges are shown as superscripts on the ions, but the charges are not shown in the formulas of the compounds). At this stage you do not need to understand the charges on these ions, but knowing about them helps to identify the numbers of atoms shown in the formulas of compounds.

Examples of polyatomic ions that form parts of compounds include:

- carbonate ( $\text{CO}_3^{2-}$ )
- sulfate ( $\text{SO}_4^{2-}$ )
- phosphate ( $\text{PO}_4^{3-}$ )
- hydroxide ( $\text{OH}^-$ )
- nitrate ( $\text{NO}_3^-$ ).

## Calculating the quantity of specific atoms from a chemical formula

An example of a compound containing polyatomic ions is aluminium sulfate, which has the formula  $\text{Al}_2(\text{SO}_4)_3$ .

### KEY TERM

**polyatomic ion** group of atoms joined together to form a charged particle

At first glance, this configuration can look very confusing; however, the brackets help you identify the polyatomic ion as a group. The subscript after the bracket—in this case the 3—applies to everything *inside* the bracket.

So, following this rule, and the rules from earlier:

- a total of two aluminium (Al) atoms and three sulfate groups (SO<sub>4</sub>).

In SO<sub>4</sub>, the subscript 4 means that each sulfate group contains one sulfur atom and four oxygen atoms, which gives:

- a total of 3 sulfur atoms (1 × 3) (one in each sulfate group)
- a total of 12 oxygen atoms (4 × 3) (four in each sulfate group).

### SC 2 CHECK YOUR UNDERSTANDING

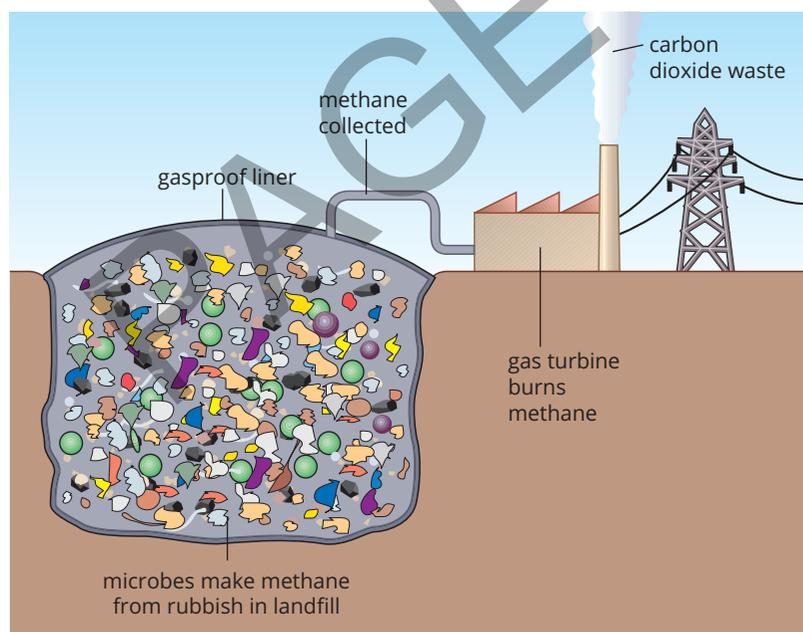
Calculate the number of atoms for each element for Ca(NO<sub>3</sub>)<sub>2</sub>.

**SC 3** I can predict the products formed in a chemical reaction based on the atomic arrangement in the reactants

## Types of chemical reactions

Knowing how atoms are rearranged in different types of chemical reactions allows you to predict the products of these reactions.

For example, in the complete combustion of methane (CH<sub>4</sub>), a carbon-based fuel, the carbon atoms in the methane combine with oxygen atoms in the air to produce carbon dioxide (CO<sub>2</sub>) while the hydrogen atoms in the methane also combine with oxygen to form water (H<sub>2</sub>O) (Figure 4.5.4).



**FIGURE 4.5.4** Left: Methane can be generated from rubbish in landfill; right: the rearrangement of atoms in the complete combustion of methane.

Table 4.5.2 shows different types of chemical reactions in more detail.

**TABLE 4.5.2** Reactants and products in different types of chemical reaction; note: a hydrocarbon contains only carbon and hydrogen

Reaction type	Key reactants	→	Key products
<b>Complete combustion</b>	hydrocarbon + excess oxygen	→	carbon dioxide + water
the hydrogen atoms in the fuel combine with oxygen in the air to form water; the carbon atoms also combine with oxygen to produce carbon dioxide			
<i>Example:</i>	$2C_4H_{10} + 13O_2$	→	$8CO_2 + 10H_2O$
<b>Incomplete combustion</b>	hydrocarbon + limited oxygen	→	carbon monoxide + water
the hydrogen atoms in the fuel combine with oxygen in the air to form water; the carbon atoms also combine with oxygen to produce carbon monoxide			
<i>Example:</i>	$2C_4H_{10} + 9O_2$	→	$8CO + 10H_2O$
<b>Acids with reactive metals</b>	acid + reactive metal	→	salt + hydrogen
the hydrogen atoms from the acid form hydrogen gas, leaving the other atoms in the acid to combine with the metal element to form a salt			
<i>Example:</i>	$Ca + H_2SO_4$	→	$CaSO_4 + H_2$
<b>Acids with bases</b>	acid + base	→	salt + water
the hydrogen atoms from the acid combine with the oxygen atoms from the base to form water, leaving the other atoms in the acid to combine with the metal element to form a salt			
<i>Example:</i>	$CaO + H_2SO_4$	→	$CaSO_4 + H_2O$
	$NaOH + HCl$	→	$NaCl + H_2O$
<b>Acids with carbonates</b>	acid + metal carbonate	→	salt + water + carbon dioxide
the hydrogen atoms from the acid combine with an oxygen atom from the carbonate ( $CO_3$ ), leaving carbon dioxide and water			
<i>Example:</i>	$CaCO_3 + H_2SO_4$	→	$CaSO_4 + H_2O + CO_2$

## Formation of salts

Remember, the word salt does not always refer to sodium chloride (table salt). Other salts can be formed, depending on the reactants. There are two parts to the name of a salt. The first word comes from the metal (either the metal or the metal in the base). The second word comes from the acid. See Figure 4.5.5 for some common reactants and their salt products.

	Base / Metal	Acid
Reactants	magnesium (Mg)	hydrochloric acid (HCl)
Name of Salt	magnesium chloride ( $MgCl_2$ )	
Reactants	sodium hydroxide (NaOH)	sulfuric acid ( $H_2SO_4$ )
Name of Salt	sodium sulfate ( $Na_2SO_4$ )	
Reactants	potassium hydroxide (KOH)	nitric acid ( $HNO_3$ )
Name of Salt	potassium nitrate ( $KNO_3$ )	
Reactants	calcium hydroxide ( $Ca(OH)_2$ )	phosphoric acid ( $H_3PO_4$ )
Name of Salt	calcium phosphate ( $Ca_3(PO_4)_2$ )	

**FIGURE 4.5.5** A simplified diagram to identify the salt product from various reactants

## SC 3 CHECK YOUR UNDERSTANDING

Predict the products formed from the reaction of hydrochloric acid (HCl) and magnesium (Mg).

## Lesson review

Use these questions to check whether you have met the learning intention for this lesson.

- 1 Define the coefficient and subscript found in chemical equations.
- 2 Name the elements present in the chemical reaction:  
 $\text{CaCO}_3 \rightarrow \text{CaO} + \text{CO}_2$
- 3 Emerald is a brilliant green gemstone.  
The chemical formula for the emerald crystal is  $\text{Be}_3\text{Al}_2(\text{SiO}_3)_6$ .  
List each element's symbol, its name (using the periodic table) and the number of atoms in the formula of emerald.
- 4 Consider the following chemical equation.  
 $\text{K}_2\text{CO}_3 + 2\text{HCl} \rightarrow 2\text{KCl} + \text{H}_2\text{O} + \text{CO}_2$ 
  - a Name the type of reaction.
  - b List all the different elements in the reaction.
- 5 Answer the following questions using the compound  $\text{FeAl}_2(\text{SO}_4)_4$ .
  - a Count the number of elements present in the molecule.
  - b Write the symbol for polyatomic ion.
  - c Calculate the number of atoms in one molecule of the compound.



## 4.6 Balancing chemical equations

### Lesson overview

Chemical reactions involve the atoms of the reactants rearranging into different combinations to form products. There are millions of different possible chemical reactions, but they all have one thing in common: the atoms that were present at the start of the reaction (in the reactants) are always still present at the end of the reaction (in the products). This is known as the law of conservation of mass. Atoms cannot be created or destroyed in a chemical reaction and anything that has mass is made up of atoms.

This means that, when you represent chemical reactions using equations, the equations must contain the same total number of each type of atom on each side of the equation. If this is not the case, the equation is said to be unbalanced, and is not following the law of conservation of mass. A balanced reaction can be thought of as a bit like a seesaw. When balanced, the total mass of the reactant atoms is the same as the total mass of the product atoms.

In this lesson, you will learn how to identify if a reaction is balanced or not. You will also learn how to construct perfectly balanced equations for some common chemical reactions.

**SC 1** I can identify whether an equation is balanced

### Introducing balanced equations

In this topic, you have learnt that when writing a chemical equation, the substances that react together (reactants) are written to the left of the arrow and the substances produced (products) are written to the right.

reactants → products

You have also learnt that the law of conservation of mass states that mass is neither created nor destroyed in chemical reactions, so the number of each type of element on each side of the equation must remain the same.

This principle is represented through a balanced chemical equation.

For example, the balanced equation of magnesium and oxygen reacting to produce magnesium oxide (Figure 4.6.1) is as follows:



In this example, two atoms of magnesium react with one molecule of oxygen, which contains two atoms, to produce magnesium oxide. The presence of the “2” coefficient in front of the magnesium oxide (2MgO) means that there are two atoms of oxygen and two atoms of magnesium on the right side of the arrow, which is necessary to keep it balanced.

### Balancing equations

Chemical equations use formulas to show the reactants and products involved in a reaction. Consider an example of how to check whether an equation is balanced according to the law of conservation of mass.

#### Learning intention

To understand how to write balanced equations for simple chemical reactions

#### Success criteria

**SC 1:** I can identify whether an equation is balanced.

**SC 2:** I can balance unbalanced equations by adding the correct coefficients.

**SC 3:** I can construct a balanced chemical equation for a range of different chemical reactions.

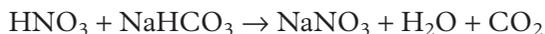


**FIGURE 4.6.1** The combustion of magnesium in air produces the white solid magnesium oxide (MgO).

## Scifile

**A cooking analogy**

Balancing a chemical equation is like making sure you have the right proportions of ingredients in a recipe. Too much or too little of one ingredient can ruin the dish.

**Neutralisation reaction between sodium hydrogen carbonate and nitric acid**

Each substance is made up of more than one type of element, and both hydrogen and oxygen are present in multiple compounds. The numbers of atoms involved in the reaction are shown in Figure 4.6.2.

Atoms	In reactants			In products				Balanced?
	HNO <sub>3</sub>	NaHCO <sub>3</sub>	Total	NaNO <sub>3</sub>	H <sub>2</sub> O	CO <sub>2</sub>	Total	
H	1	1	2	0	2	0	2	✓
N	1	0	1	1	0	0	1	✓
O	3	3	6	3	1	2	6	✓
Na	0	1	1	1	0	0	1	✓
C	0	1	1	0	0	1	1	✓

**FIGURE 4.6.2** The equation is balanced because all of the atoms present in the reactants are also present in the products.

**SC 1 CHECK YOUR UNDERSTANDING**

Explain why the following equation is not balanced.

**SC 2**

I can balance unbalanced equations by adding the correct coefficients

**Balancing equations with coefficients**

In chemical equations, coefficients and subscripts are not the same. Coefficients state how many molecules or atoms of substance or compound are involved in the reaction, whereas subscripts represent the number of atoms or ions present in each formula.

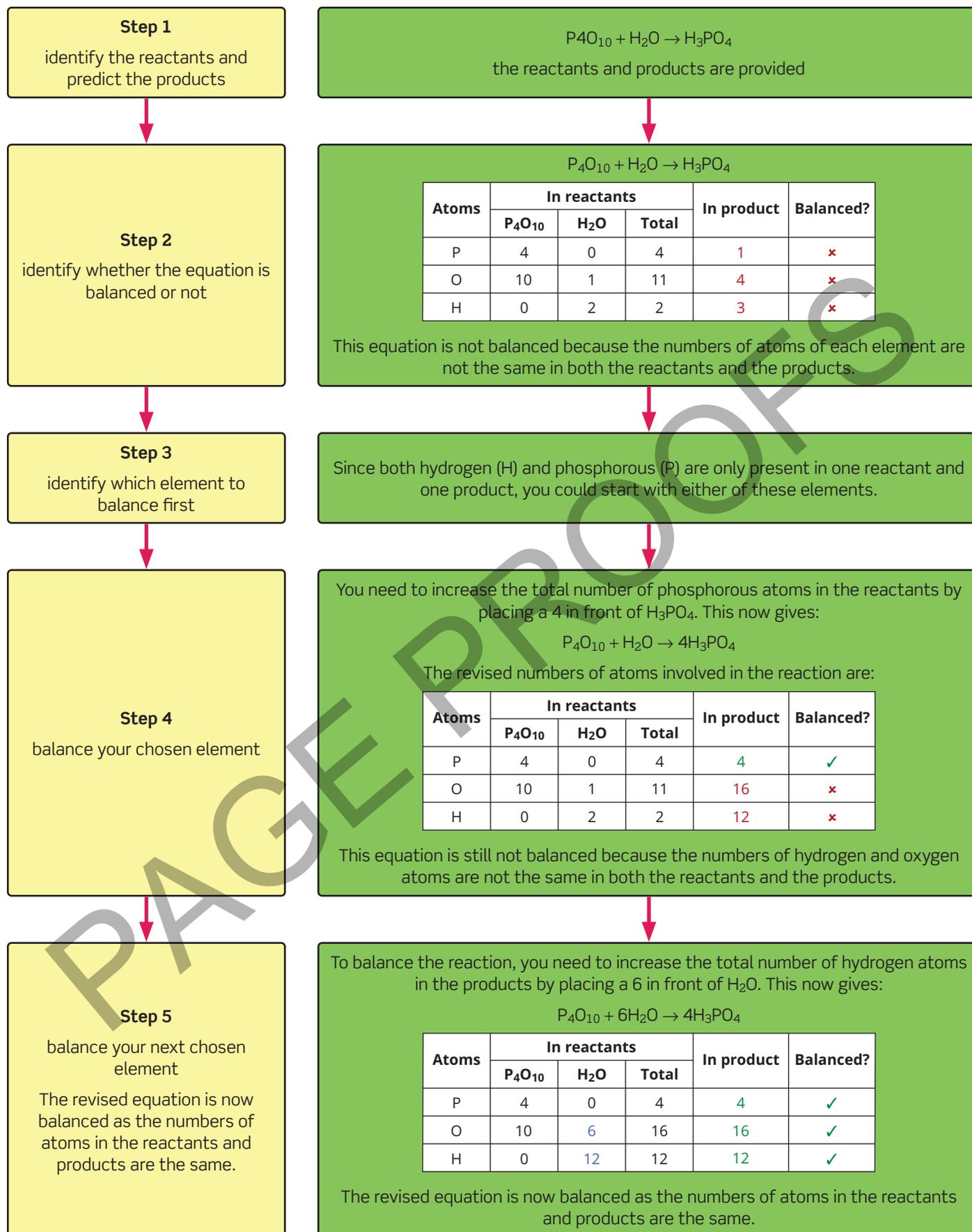
For example:



SO<sub>2</sub> and SO<sub>3</sub> are two different chemicals (sulfur dioxide and sulfur trioxide respectively).

Their subscripts are critical to understanding the identity and properties of each compound. The coefficient '2' in front of both molecules indicates that, to balance the number of atoms present in the equation, you require two molecules of sulfur dioxide and two molecules of sulfur trioxide, but only one oxygen molecule.

This means that you can only balance a chemical reaction by changing the coefficients—not by changing the subscripts. Changing the subscripts would change the substances, making the equation incorrect. Follow the steps in Figure 4.6.3, which includes the formation from phosphorus pentoxide and water as an example.



**FIGURE 4.6.3** Step guide to balancing chemical equations, using the formation from phosphorus pentoxide and water as an example.

Any unbalanced chemical equation can be balanced using this process. You do not need to write out a table for every problem, but you can use the same thought processes as you work through each type of atom.

### SC 2 CHECK YOUR UNDERSTANDING

Balance the following equation.



**SC 3** I can construct a balanced chemical equation for a range of different chemical reactions

### HINT

Just counting the nitrate 'groups' makes it easier to balance the equation, rather than counting the individual nitrogen and oxygen atoms. In this case, the aim is to get two nitrate groups on each side of the equation.

### Representing types of chemical reactions

In this topic you have encountered four specific types of reactions: combustion, neutralisation, metal–acid and corrosion reactions.

It is important to understand that whatever the reaction type, the process of balancing the chemical equations is the same. You can use the reaction type to predict the products, then create a balanced chemical equation using the formulas of the reactants and products involved in the reaction. Figure 4.6.4 has the steps for balancing a neutralisation reaction and Figure 4.6.5 has the steps for balancing a combustion reaction.

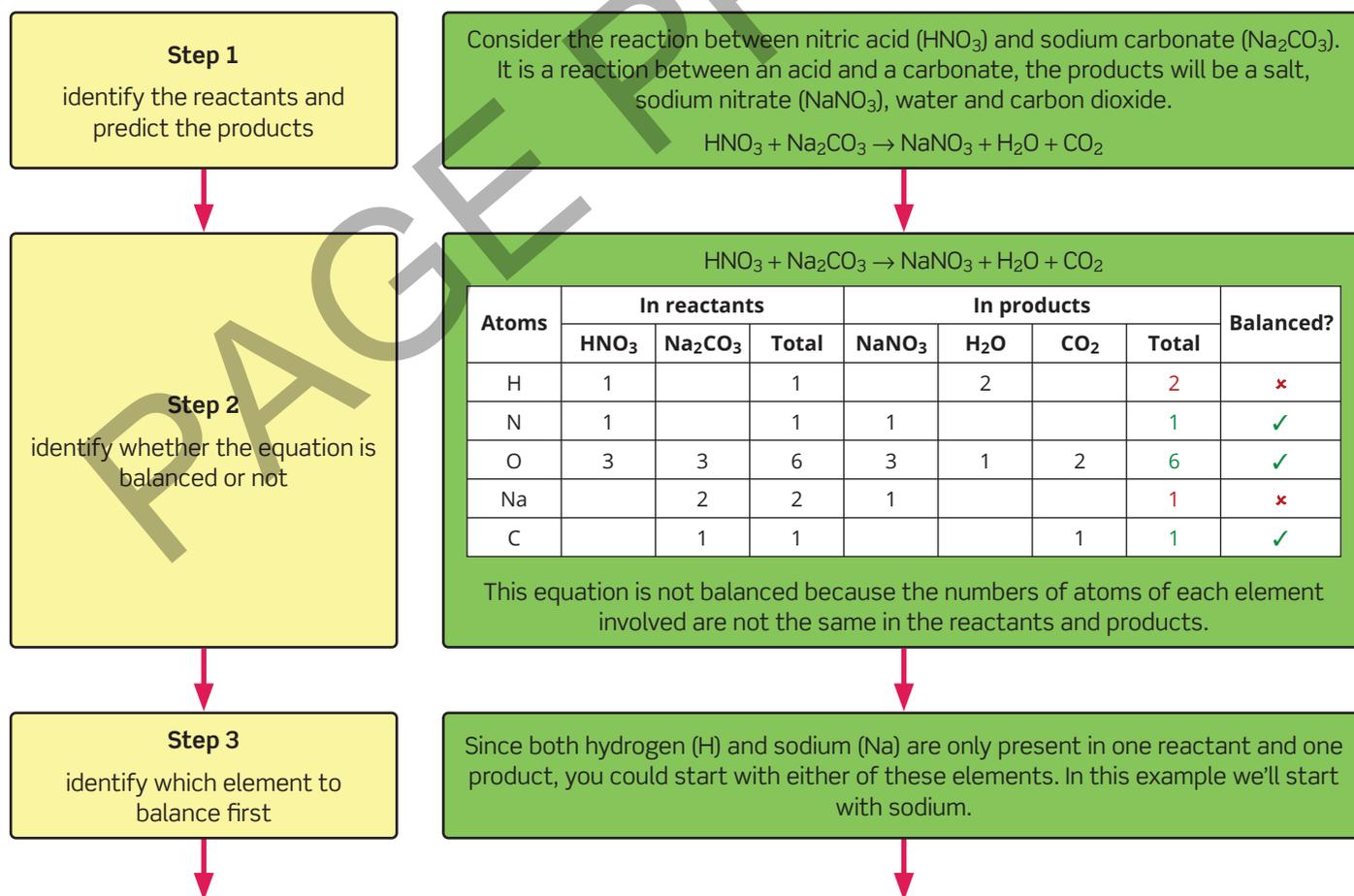


FIGURE 4.6.4 (continued)

**Step 4**

balance your chosen element

First, you can increase the number of sodium atoms on the right side of the arrow. To do this, the coefficient 2 can be placed in front of  $\text{NaNO}_3$ . This now gives:

$$\text{HNO}_3 + \text{Na}_2\text{CO}_3 \rightarrow 2\text{NaNO}_3 + \text{H}_2\text{O} + \text{CO}_2$$

Atoms	In reactants			In products				Balanced?
	$\text{HNO}_3$	$\text{Na}_2\text{CO}_3$	Total	$2\text{NaNO}_3$	$\text{H}_2\text{O}$	$\text{CO}_2$	Total	
H	1		1		2		2	✗
N	1		1	2			2	✗
O	3	3	6	6	1	2	9	✗
Na		2	2	2			2	✓
C		1	1			1	1	✓

However, this equation is still not balanced because the numbers of hydrogen, nitrogen and oxygen atoms are not the same in the products and reactants.

**Step 5**

balance your next chosen element

The revised equation is now balanced as the numbers of atoms in the reactants and products are the same.

To balance the equation, you need to increase the number of hydrogen, nitrogen and oxygen atoms on the left side of the arrow. To do this, the coefficient 2 can be placed in front of  $\text{HNO}_3$ . This now gives:

$$2\text{HNO}_3 + \text{Na}_2\text{CO}_3 \rightarrow 2\text{NaNO}_3 + \text{H}_2\text{O} + \text{CO}_2$$

Atoms	In reactants			In products				Balanced?
	$2\text{HNO}_3$	$\text{Na}_2\text{CO}_3$	Total	$2\text{NaNO}_3$	$\text{H}_2\text{O}$	$\text{CO}_2$	Total	
H	2		2		2		2	✓
N	2		2	2			2	✓
O	6	3	9	6	1	2	9	✓
Na		2	2	2			2	✓
C		1	1			1	1	✓

The equation is now balanced.

**FIGURE 4.6.4** Steps to balance a neutralisation equation

**Step 1**

identify the reactants and predict the products

Consider the reaction between the hydrocarbon pentene ( $\text{C}_5\text{H}_{10}$ ) and oxygen. Pentene and oxygen are the reactants and so are written to the left of the arrow in a chemical equation. As the reaction is performed in oxygen – that is, the combustion is complete – the products will be carbon dioxide ( $\text{CO}_2$ ) and water ( $\text{H}_2\text{O}$ ):

$$\text{C}_5\text{H}_{10} + \text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O}$$

**Step 2**

identify whether the equation is balanced or not

$\text{C}_5\text{H}_{10} + \text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O}$

Atoms	In reactants			In products			Balanced?
	$\text{C}_5\text{H}_{10}$	$\text{O}_2$	Total	$\text{CO}_2$	$\text{H}_2\text{O}$	Total	
C	5	0	5	1	0	1	✗
H	10	0	10	0	2	2	✗
O	0	2	2	2	1	3	✗

This equation is not balanced because the numbers of atoms of each element involved are not the same in the reactants and products.

**FIGURE 4.6.5** (continued)

**Step 3**  
identify which element to balance first

Since both hydrogen (H) and carbon (C) are only present in one reactant and one product, you could start with either of these elements.

**Step 4**  
balance your chosen element

First, you need to increase the total number of carbon atoms in the products by placing the coefficient 5 in front of CO<sub>2</sub>. This now gives:

$$\text{C}_5\text{H}_{10} + \text{O}_2 \rightarrow 5\text{CO}_2 + \text{H}_2\text{O}$$

Atoms	In reactants			In products			Balanced?
	C <sub>5</sub> H <sub>10</sub>	O <sub>2</sub>	Total	CO <sub>2</sub>	H <sub>2</sub> O	Total	
C	5	0	5	5	0	5	✓
H	10	0	10	0	2	2	✗
O	0	2	2	10	1	11	✗

However, this equation is still not balanced because the numbers of hydrogen and oxygen atoms are not the same in both the reactants and products.

**Step 5**  
balance your next chosen element

The revised equation is now balanced as the numbers of atoms in the reactants and products are the same.

Next, you need to increase the total number of hydrogen atoms in the products by placing the coefficient 5 in front of H<sub>2</sub>O. This now gives:

$$\text{C}_5\text{H}_{10} + \text{O}_2 \rightarrow 5\text{CO}_2 + 5\text{H}_2\text{O}$$

Atoms	In reactants			In products			Balanced?
	C <sub>5</sub> H <sub>10</sub>	O <sub>2</sub>	Total	CO <sub>2</sub>	H <sub>2</sub> O	Total	
C	5	0	5	5	0	5	✓
H	10	0	10	0	10	10	✓
O	0	2	2	10	5	15	✗

However, this equation is still not balanced because the numbers of oxygen atoms are not the same in both the reactants and products.

To balance the equation, you need to increase the total number of oxygen atoms in the products by placing the coefficient 7.5 in front of O<sub>2</sub>. This now gives:

$$\text{C}_5\text{H}_{10} + 7.5\text{O}_2 \rightarrow 5\text{CO}_2 + 5\text{H}_2\text{O}$$

Atoms	In reactants			In products			Balanced?
	C <sub>5</sub> H <sub>10</sub>	O <sub>2</sub>	Total	CO <sub>2</sub>	H <sub>2</sub> O	Total	
C	5	0	5	5	0	5	✓
H	10	0	10	0	10	10	✓
O	0	15	15	10	5	15	✓

The equation is now balanced. However, balanced equations cannot contain fractions of molecules, i.e. 7.5 oxygen molecules.

To correct this problem, double each of the coefficients. The final balanced chemical equation is therefore:

$$2\text{C}_5\text{H}_{10} + 15\text{O}_2 \rightarrow 10\text{CO}_2 + 10\text{H}_2\text{O}$$

FIGURE 4.6.5 Steps to balance a combustion equation

## HINT

In combustion reactions, always leave oxygen to balance last as it is present as an element (in the reactants). You can follow this approach to create balanced equations for the many types of reactions that you will encounter in Year 9.

### SC 3 CHECK YOUR UNDERSTANDING

Write out and balance the incomplete combustion reaction of butane ( $C_4H_{10}$ ) to carbon monoxide.

## Lesson review

Use these questions to check whether you have met the learning intention for this lesson.

- Identify the following chemical reactions as balanced or unbalanced.
  - $SnO_2 + H_2 \rightarrow Sn + H_2O$
  - $N_2 + O_2 \rightarrow 2NO$
  - $2CH_4 + 2O_2 \rightarrow 2CO + 4H_2O$
  - $CuO + H_2SO_4 \rightarrow CuSO_4 + H_2O$
  - $2Al + 2HCl \rightarrow 2AlCl_3 + H_2$
- Balance the chemical reactions in question 1 that are unbalanced.
- Calculate the value of the coefficient 'x' in the following chemical reactions.
  - $TiCl_4 + xH_2O \rightarrow TiO_2 + 4HCl$
  - $2H_2 + xO_2 \rightarrow 2H_2O$
  - $3CaCl_2 + 2Na_3PO_4 \rightarrow Ca_3(PO_4)_2 + xNaCl$
- Write the balanced chemical equation for the neutralisation reaction between potassium hydroxide KOH and phosphoric acid  $H_3PO_4$ .

## 4.7 Advantages of green chemistry

### Learning intention

To be able to evaluate why green chemistry processes are being adopted by manufacturers

### Success criteria

**SC 1:** I can explore examples of green chemistry in manufacturing.

**SC 2:** I can evaluate why manufacturers use green chemistry.

**SC 3:** I can construct a representation detailing the advantages of green chemistry for society.



**FIGURE 4.7.1** Green chemistry principles focus on creating chemical processes that are less harmful, use less energy, or produce less waste; biodiesel is an example of a product that is less harmful, uses a food waste product (fats and oils from food industry) and does not require fossil fuel extraction.

### KEY TERMS

**solvent** a substance that dissolves another substance  
**emission** the release of light, heat or gas

## Introduction

In response to growing worries about the environment and the need to reduce the negative effects of industry, manufacturers are increasingly adopting green chemistry processes. Traditional manufacturing methods cause pollution and use up valuable resources. Green chemistry, by contrast, focuses on creating chemical processes that are less harmful, use less energy, or produce less waste (Figure 4.7.1). By using green chemistry principles, manufacturers hope to make their operations more sustainable and save money by increasing efficiency and reducing the costs associated with following regulations.

In this inquiry activity, you will evaluate why manufacturers are adopting green chemistry processes by reflecting on the benefits it can offer them.

## Background

Green chemistry is a broad term that covers a wide range of processes used in the chemical industry. It includes producing less waste material, using safer chemicals in manufacturing processes, reducing energy requirements and being more efficient in the use of Earth's resources.

Industrial manufacturers, guided by chemical engineers, can design processes that incorporate the key principles of green chemistry. However, industry decisions are based on many things, including profit margins, availability of staffing, access to resources and current market forces. Accordingly, there are many factors that influence how 'green' the chemical industry is.

As the scope of green chemistry is so wide, it is suggested that you focus on one or two principles in this inquiry, and research them in detail in order gain enough information to represent your ideas.

## Aim

To evaluate why green chemistry processes are being adopted by some manufacturers

## Plan

Conduct some research to investigate green chemistry processes in action. Try to find examples in Australia, if possible in the region where you live, and outline why you think manufacturers are embracing them. Below are some examples of the application of green chemistry to give ideas for your research.

### Solvent substitution

Manufacturers are opting to use less-hazardous **solvents** or solvent-free processes, minimising harmful **emissions** and waste production.

## Catalysis

Green catalysts are used to enhance chemical reactions, reducing energy consumption and waste generation while increasing efficiency.

## Biodegradable polymers

The production of biodegradable plastics and polymers reduces the persistence of plastic waste in the environment.

## Renewable feedstocks

Using renewable resources, like plant-based raw materials, decreases dependence on fossil fuels and reduces **carbon footprint**.

## Microwave and ultrasonic assisted reactions

These technologies accelerate reactions, leading to shorter reaction times, decreased energy usage and increased efficiency. For example, a traditional process that produces a product could take 24 hours whereas the microwave assisted process could reduce the time to 90 minutes.

## Continuous flow processes

Implementing continuous manufacturing processes reduces the need for large batch reactions, minimising waste and energy consumption.

## Waste minimisation

Green chemistry encourages the design of processes that generate fewer by-products and waste, thus minimising environmental impact.

## Water-based processes

Substituting water for hazardous solvents in reactions reduces environmental contamination and risk.

## Life cycle analysis

Manufacturers analyse the entire life cycle of a product to identify areas for improvement, such as reducing energy consumption during production and disposal.

## Green analytical techniques

Employing environmentally friendly analytical methods minimises the use of toxic reagents (substances used in chemical analysis) and reduces waste generation during testing.

## Design

- 1 This inquiry activity requires you to create a representation that explores and outlines the advantages of green chemistry for society. Using the information you gathered on the green chemistry principle of your choice, record the key headings and details that you would include in your representation.

### KEY TERM

**carbon footprint** a measure of the total amount of the greenhouse gases carbon dioxide and methane produced by the actions of an individual, population or industry

- 2** You can now decide on the format for your representation. This could be in the form of a:
- mind map that links the various ideas around green chemistry
  - SWOT (strengths, weaknesses, opportunities and threats) analysis
  - newspaper article about a particular case study
  - quiz based on green chemistry
  - persuasive text, such as a letter to a chemical manufacturer from an environmental agency encouraging green energy approaches.

Whatever format you choose, select and incorporate information from your research, including local examples where possible. Seek advice from your teacher as the audience the representation is aimed at.

## Conduct

Using all the information in your planning and design, construct your representation of the advantages of green chemistry for manufacturing and for society. Include examples where you can and add images if you want to illustrate your ideas.

## Improve

Consider how you might improve your representation. What information was missing or could have been changed?

## Evaluate

- 1** What aspect of chemical reactions was explored through this inquiry into green chemistry?
- 2** What skills did you use in this inquiry?

## 4.8 Combustion reactions of fuels

### Introduction

For more than a century, fossil fuels have been used in combustion reactions to provide energy for society. Because the compounds in fossil fuels contain the elements hydrogen and carbon, they are often referred to as hydrocarbons. In a combustion reaction, these carbon-based fuels react with oxygen, producing new products and releasing heat (Figure 4.8.1). The release of heat energy means that it is an exothermic reaction.

In this practical investigation, you will explore the combustion of different fuels to identify the products of these reactions.

### Background

Complete combustion occurs when there is plenty of oxygen to react with the carbon-based fuel. The reaction can therefore use as much oxygen as it needs to form carbon dioxide as one of the products.

The general equation for complete combustion is:

carbon-based fuel + oxygen  $\rightarrow$  carbon dioxide + water

If the amount of oxygen available to react is restricted, incomplete combustion occurs and carbon monoxide, a toxic gas, is formed instead of the carbon dioxide.

The general equation for incomplete combustion is:

carbon-based fuel + oxygen  $\rightarrow$  carbon monoxide + water

You will be testing the products of combustion fuels containing hydrogen and carbon to see whether the expected products were produced. Limewater is used to test for the presence of carbon dioxide. Cobalt chloride paper is used to test water production.

### Aim

To identify the products of combustion of different carbon-based fuels

### Materials

- glass funnel larger than 6 cm in diameter
- boiling tube containing a piece of blue cobalt chloride paper sealed with a two-holed rubber bung, one hole fitted with a long piece of glass tubing and the other hole fitted with a short piece of glass tubing (see Figure 4.8.2)
- boiling tube containing about 20 mL of limewater sealed with a two-holed rubber bung, one hole fitted with a long piece of glass tubing and the other hole fitted with a short piece of glass tubing (see Figure 4.8.2)
- rubber or silicone tubing for connections
- filtering pump
- spirit burner with ethanol and propanol
- forceps
- molecular models or materials that can be used to create models of molecules
- matches

### Learning intention

To be able to observe combustion reactions and model the atomic rearrangements occurring during the reactions

### Success criteria

**SC 1:** I can safely carry out and record observations of combustion reactions of carbon-based fuels.

**SC 2:** I can use models to represent atomic rearrangement in combustion reactions of simple hydrocarbons.

**SC 3:** I can create and use balanced chemical equations to represent combustion reactions of simple hydrocarbons.



**FIGURE 4.8.1** A household 'bio fire' that uses the combustion of ethanol, a carbon-based fuel, to create the natural flame.

## Assessment of risk

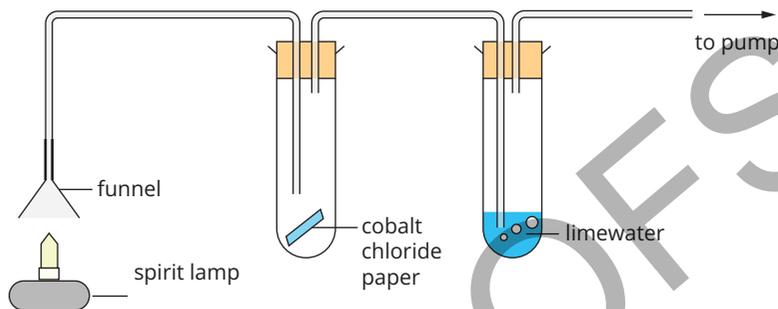
Ensure you are aware of the risks of this practical investigation and have considered how safety can be improved before carrying out this activity.

### SAFETY NOTES

- ▶ Be careful around open flames.
- ▶ Be careful with glassware to avoid breaking it. If glassware is broken, follow the teacher's instructions and take care as sharp shards of glass could remain on the bench.
- ▶ Do not connect plastic tubing to the glass funnel; it will get very hot and can melt.
- ▶ Use forceps to place the cobalt chloride paper into the test tube. Do not handle the paper.
- ▶ Once the experiment is finished, allow the equipment to cool so that it is safe to handle.
- ▶ Only test for ethanol and propanol. Do not use other liquid fuels.

## Method

- 1 Set up the apparatus as shown in Figure 4.8.2. Ensure that the connections to the boiling tubes are the correct way round and the tube entering the second boiling tube reaches below the surface of the limewater.



**FIGURE 4.8.2** Simplified diagram of the experimental setup

- 2 Switch on the pump to draw a gentle stream of air through the apparatus.
- 3 Light the spirit lamp filled with the liquid fuel.
- 4 Wait for a few minutes to see if the cobalt chloride paper turns from blue to pink, indicating the presence of water, and if the limewater goes milky (it produces a white precipitate), indicating the presence of carbon dioxide.

## Results

Record your results in a table like the one below.

Fuel	Test performed	Observations
	limewater	
	cobalt chloride paper	

## Conclusion

- 1 State the products that can be identified using limewater and cobalt chloride paper. Using this information, write a word equation for each of the reactions investigated.
- 2 Use commercial or home-made playdough, molecular models or computer simulations to model the rearrangement of atoms in the complete combustion of ethane ( $C_2H_6$ ).
- 3 Write a balanced equation for the complete combustion of one of the compounds burned in the experiment.
- 4 Write a conclusion for this investigation that includes an overview of the experiment and your key findings.

## Evaluation

This practical investigation established that the combustion of different hydrocarbons produced carbon dioxide and water. However, it did not test for carbon monoxide and so there is no way to determine whether combustion was fully complete. Furthermore, time only permitted for a few compounds to be studied.

Complete an evaluation of this experiment and propose how it could be modified to allow for a more thorough investigation of whether the general formulas for combustion reactions hold true.

### HINT

If you need help with balancing chemical equations, refer to previous lessons.

PAGE PROOFS

# Chemical reactions: Rearranging atoms

## Topic summary

The key concepts included in this topic are:

- Atoms are not created or destroyed in chemical reactions; they are rearranged to form new products.
- The law of conservation of mass states that mass is conserved in a chemical reaction, and this is represented by using balanced chemical equations.
- Neutralisation reactions form salts and can be used to reduce acidity.
- The products from the reactions of acids can be predicted from the name of the acid and the nature of the other reactant.
- Corrosion reactions of metals involve the reaction with gases present in the atmosphere.
- Combustion reactions require oxygen and are exothermic, giving of heat and light, and producing oxides as products. The nature of the products can depend on the amount of oxygen available.
- The application of the principles of green chemistry improves sustainability and includes reducing the use of toxic substances, using less energy, following more efficient processes, creating less waste and using renewable resources.

## Review questions

The following questions will assess your success in achieving the learning intentions for this topic.

### Remember

- 1 Name the salt that is produced when sulfuric acid reacts with barium.
- 2 Determine the mass of products you would expect when the total mass of reactants is 4.492 g.
- 3 Recall the main environmental factor that affects combustion.

### Understand

- 4 Explain the difference between complete combustion and incomplete combustion.
- 5 Name the type of reaction between hydrochloric acid and baking soda ( $\text{NaHCO}_3$ ).
- 6 Describe how a manufacturer might benefit from using a green chemistry process to produce a product.

### Apply

- 7 Explain why a new copper dome on a cathedral turns green so after it is installed.

- 8 Explain why rust forms more quickly on boats on the sea compared to boats on a freshwater lake.
- 9 Write out the products of the following chemical reactions:
  - a sodium hydroxide + hydrochloric acid
  - b butane + limited oxygen
  - c nitric acid + calcium carbonate
  - d silver + hydrogen sulfide

### Analyse

- 10 Analyse the potential sources of error when measuring the mass of a reactant in a chemical reaction.
- 11 Consider the reaction between aluminium Al and hydrochloric acid HCl. The products are aluminium chloride  $\text{AlCl}_3$  and hydrogen gas  $\text{H}_2$ .
  - a Write the balanced chemical equation for the reaction.
  - b Explain the atomic rearrangement that occurs.
  - c Discuss how the conservation of mass is demonstrated in this reaction.

- 12** Analyse the atomic rearrangement in the combustion of ethanol  $\text{C}_2\text{H}_5\text{OH}$ .
- 13** Balance the following chemical equations:
- a**  $\text{AgI} + \text{Na}_2\text{S} \rightarrow \text{Ag}_2\text{S} + \text{NaI}$
  - b**  $\text{PCl}_5 + \text{H}_2\text{O} \rightarrow \text{H}_3\text{PO}_4 + \text{HCl}$
  - c**  $\text{Mg}_3\text{N}_2 + \text{H}_2\text{O} \rightarrow \text{MgO} + \text{NH}_3$
  - d**  $\text{Fe}_3\text{O}_4 + \text{CO} \rightarrow \text{FeO} + \text{CO}_2$
  - e**  $\text{NH}_4\text{Cl} + \text{NaNO}_2 \rightarrow \text{N}_2 + \text{H}_2\text{O} + \text{NaCl}$
  - f**  $\text{N}_2 + \text{H}_2 \rightarrow \text{NH}_3$
  - g**  $\text{Li} + \text{H}_2\text{O} \rightarrow \text{LiOH} + \text{H}_2$
  - h**  $\text{C}_4\text{H}_{10} + \text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O}$
  - i**  $\text{C}_4\text{H}_{10} + \text{O}_2 \rightarrow \text{CO} + \text{H}_2\text{O}$

## Extension

- 14** Combustion and photosynthesis are both chemical reactions involving energy changes and atomic rearrangements.

Research and compare the similarities and differences between combustion reactions and photosynthesis.

Present your findings in a 5-minute presentation, highlighting the key components, atomic rearrangements, and energy changes involved in each process. Include examples of how these reactions are essential for life and industry. You could use a diagram of photosynthesis and a combustion reaction to support your explanation.

## Topic reflection

The learning intentions for this topic are given in each lesson and at the beginning of the topic. Consider how well you have achieved them. Note down any particular areas that you are confident in, and others where you are not so sure.

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## 4

## Glossary

**balanced equation** a chemical equation that has the same number of each type of atom on both sides of the equation

**base** a substance that will neutralise an acid

**carbon footprint** a measure of the total amount of the greenhouse gases carbon dioxide and methane produced by the actions of an individual, population or industry

**chemical equation** a short-hand notation that scientists use to communicate what happens during a chemical reaction

**chemical formula** a representation that uses symbols to indicate the elements in a substance and the relative

**coefficient** the big number in front of a chemical in a balanced equation

**combustion** a chemical reaction in which a substance burns in oxygen gas to produce light and heat

**complete combustion** that occurs when there is plenty of oxygen available; it produces carbon dioxide and water vapour

**corrosion** a chemical reaction in which a metal reacts with oxygen to produce a metal oxide but does not produce significant amounts of heat and light

**desiccant** a substance that absorbs water from the surrounding air

**electrolyte** a substance that can dissolve to form a solution that conducts electricity and/or can conduct electricity in its molten form

**hydrated** a substance that contains water

**incomplete combustion** combustion that occurs when oxygen is limited; produces carbon (soot, smoke) and carbon monoxide, and does not release as much heat or light as complete combustion

**law of conservation of mass** the law that states that atoms cannot be created or destroyed during a chemical reaction

**neutralisation** a reaction of an acid with a base, forming a salt and water

**non-renewable** source of energy that cannot be easily replaced after it is used, such as fossil fuels like coal or gas

**polyatomic ion** group of atoms joined together to form a charged particle

**polymer** a very large molecule composed of many repeating subunits

**precipitate** a solid formed during a chemical reaction

**precipitation reaction** when two clear solutions react to produce an insoluble solid

**product** a substance produced by a chemical reaction

**reactant** a substance that takes part in a chemical reaction

**rust** hydrated iron(III) oxide; chemical formula  $\text{Fe}_2\text{O}_3$

**solvent** a substance that dissolves another substance

**tarnish** a black coating of silver sulfide that is produced when silver reacts with sulfur in food or the atmosphere; chemical formula  $\text{Ag}_2\text{S}$

**verdigris** a bluish-green layer that forms on copper, brass or bronze due to reaction with oxygen, water and carbon dioxide

**word equation** a chemical equation in which the reactants and products are identified by their chemical names

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