Chemical reactions

5

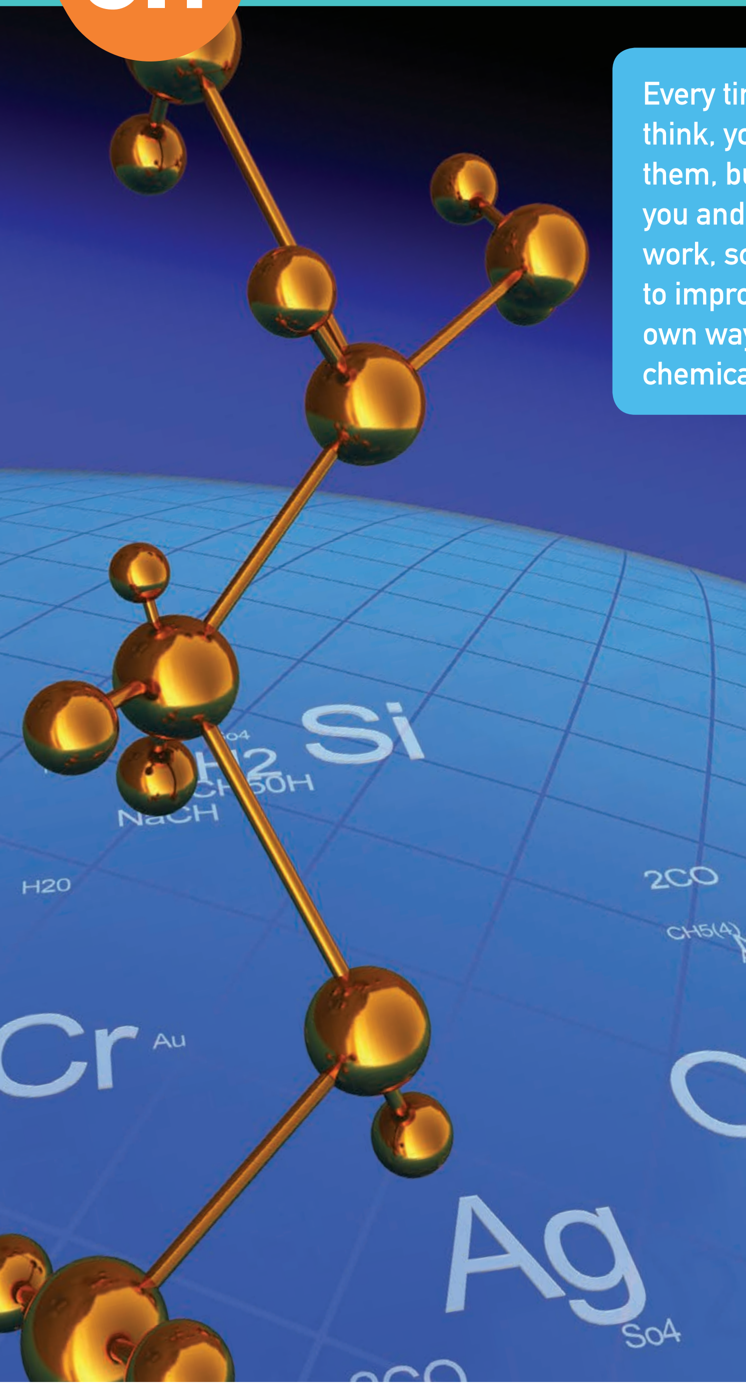
HAVE YOU EVER WONDERED…

- what makes a bullet explode?
- how a candle stays alight?
- where a battery gets its energy from?
- why you put milk in the fridge?

After completing this chapter students should be able to:

- use word and formula equations to represent chemical reactions
- predict the products of different types of simple chemical reactions
- describe how chemistry can be used to produce a range of useful substances such as fuels, metals and pharmaceuticals
- explain the effect of a range of factors including temperature, surface area, concentration, agitation and catalysts on the rate of chemical reactions.

5.1 Chemical equations



Every time you light a candle, bake a cake, take a breath or even think, you start a chemical reaction. You may not always notice them, but chemical reactions are going on constantly around you and inside you. By understanding how chemical reactions work, scientists are able to use and control chemical reactions to improve our quality of life. Chemists have developed their own way of describing and explaining what happens during a chemical reaction. They do this by writing chemical equations.

What is a chemical equation?

Chemical equations are one of the most basic tools for describing what happens during a chemical reaction. A chemical equation can communicate detailed information about any chemical reaction in a single line. The general structure of a chemical equation is:

reactants → products

where the arrow (→) means 'rearrange to form'.

Word and formula equations

Replacing the **reactants** and **products** with their chemical names gives a **word equation**. For example, when calcium carbonate ($CaCO_3$) reacts with sulfuric acid (H_2SO_4), it produces calcium sulfate ($CaSO_4$), water (H_2O) and carbon dioxide gas (CO_2). In this reaction, you have:

Reactants = calcium carbonate, sulfuric acid

Products = calcium sulfate, water, carbon dioxide gas

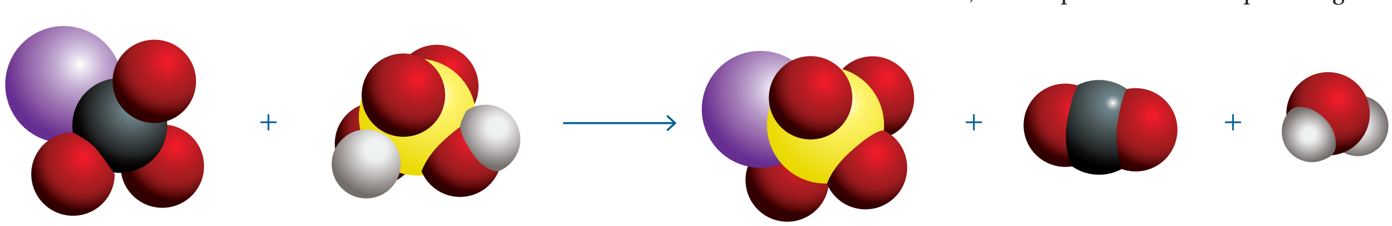
So the general form of an equation:

reactants → products

becomes the word equation:

calcium carbonate + sulfuric acid → calcium sulfate + carbon dioxide gas + water

This is much simpler than trying to explain the chemical reaction with sentences. However, word equations are still quite long.



calcium carbonate + sulfuric acid → calcium sulfate + carbon dioxide + water

$CaCO_3 + H_2SO_4 \rightarrow CaSO_4 + CO_2 + H_2O$

Figure 5.1.1 The atoms from the calcium carbonate combine with the atoms from the sulfuric acid to produce calcium sulfate, hydrogen and carbon dioxide.

Scientists can write the equation much more effectively by replacing the chemical names with their chemical formulas. The chemical equation for the previous example then becomes:

$$CaCO_3 + H_2SO_4 \rightarrow CaSO_4 + CO_2 + H_2O$$

This type of chemical equation is known as a **formula equation**. The formula equation is shorter to write and contains more information. It shows exactly what atoms (or ions) are involved in the chemical reaction. Figure 5.1.1 shows how the reactants in the equation above rearrange to form the products.

Showing states

The reactants and products in a chemical reaction can be in one of four states—solid, liquid, gas or **aqueous solution** (dissolved in water). Chemists give these states the symbols:

- (s) for a solid
- (l) for a liquid
- (g) for a gas
- (aq) for an aqueous solution.

The symbols can be added to the formula equations to give even more information about the chemical reaction. In the previous example, calcium carbonate and calcium sulfate are both solids, sulfuric acid is an aqueous solution, water is a liquid and carbon dioxide is a gas. All of this information can be included in the formula equation by writing the symbol for the state next to the formula name:

$$CaCO_3(s) + H_2SO_4(aq) \rightarrow CaSO_4(s) + H_2O(l) + CO_2(g)$$

The states are always written after the formula name in brackets.

Balanced chemical equations

The formula equation for the reaction of calcium carbonate with sulfuric acid is a **balanced equation**. This means that it has the same number of atoms of each element on both sides of the equation. You can easily check this by counting the number of atoms in the reactants and products:

Reactants = $1 \times Ca, 1 \times C, 7 \times O, 2 \times H, 1 \times S$

Products = $1 \times Ca, 1 \times C, 7 \times O, 2 \times H, 1 \times S$

Balanced equations are consistent with the **law of conservation of mass**. This law states that:

During a chemical reaction, atoms cannot be created or destroyed.

You cannot create or destroy atoms in a chemical reaction. But you can rearrange them. As a result, the number of atoms in the reactants must equal the number of atoms in the products. Also, the mass of the reactants must equal the mass of the products. This is shown in Figure 5.1.2.

Figure 5.1.2 The total mass of these beakers does not change despite a chemical reaction having taken place. The law of conservation of mass says that the total mass of the reactants and the total mass of the products must be the same.

However, not all formula equations will be balanced when you first write them. For example, when hydrogen gas reacts with oxygen gas, the product is water. The word and formula equations for this reaction are:

$$\text{hydrogen} + \text{oxygen} \rightarrow \text{water}$$

$$H_2(g) + O_2(g) \rightarrow H_2O(l)$$

Counting the number of atoms on both sides of the equation shows that the equation is not balanced.

Reactants = $2 \times H, 2 \times O$

Products = $2 \times H, 1 \times O$

This means that if one molecule of hydrogen reacts with one molecule of oxygen, then an oxygen atom is left over. This is shown in Figure 5.1.3.

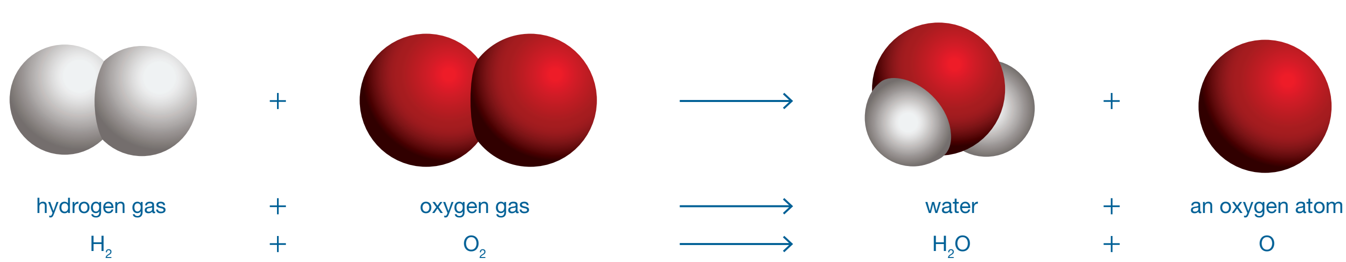
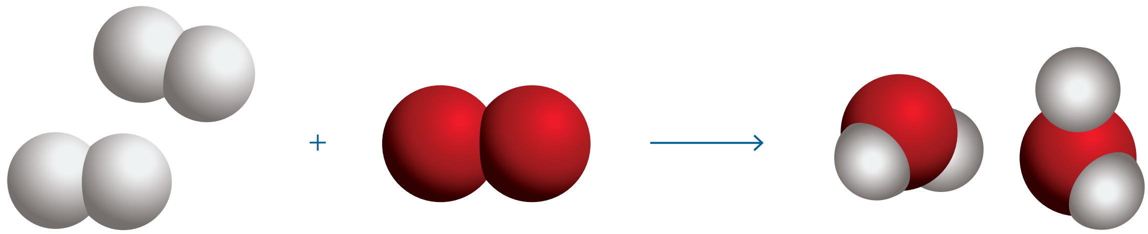


Figure 5.1.3 A single hydrogen molecule will not react with a single oxygen molecule because there would be an extra oxygen atom left over.



hydrogen gas + oxygen gas → water

$2H_2 + O_2 \rightarrow 2H_2O$

Figure 5.1.4 It takes two hydrogen molecules for each oxygen molecule, and this produces two water molecules.

However, if two hydrogen molecules react with one oxygen molecule, then the atoms can rearrange to produce two complete molecules of water. This is shown in Figure 5.1.4. Chemists represent this reaction as a balanced formula equation by writing:

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$$

Placing a 2 in front of the chemical formula for hydrogen and water indicates that the reaction uses two hydrogen molecules and produces two water molecules. Re-counting the number of atoms in the reactants and products shows that this equation is now balanced.

Reactants = 4 × H, 2 × O

Products = 4 × H, 2 × O

Consider another chemical reaction in which calcium metal (Ca) reacts with oxygen gas (O_2) to produce solid calcium oxide (CaO). The reactants in this reaction are calcium and oxygen gas. The only product is calcium oxide. Therefore, the general equation:

$$\text{reactants} \rightarrow \text{products}$$

becomes the word equation:

$$\text{calcium} + \text{oxygen} \rightarrow \text{calcium oxide}$$

Replacing the chemical names with their formulas gives the formula equation:

$$Ca + O_2 \rightarrow CaO$$

Adding the states to the formula equation gives:

$$Ca(s) + O_2(g) \rightarrow CaO(s)$$

Checking the atoms of each element on both sides shows that the equation is unbalanced:

Reactants = 1 × Ca, 2 × O

Products = 1 × Ca, 1 × O

However, it will be balanced if two calcium atoms react with one oxygen molecule to produce two CaO molecules. So the final balanced equation can be written as:

$$2Ca(s) + O_2(g) \rightarrow 2CaO(s)$$

SciFile

Big burners

The fuel tanks that supply fuel to power space rockets into orbit do not use petrol. Some use oxygen and hydrogen that react explosively to produce large amounts of energy.

The equation is:

$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$

This creates the enormous plume of exhaust gases that exit from the rocket motors. The rockets accelerate the massive weight to reach escape velocity of about 11 km/s.

SKILL BUILDER

Balancing equations

Balancing chemical equations can be tricky, but if you follow some simple steps you should arrive at the right answer. Let's look at the reaction between potassium carbonate (K_2CO_3) and nitric acid (HNO_3), which produces potassium nitrate (KNO_3), water (H_2O) and carbon dioxide (CO_2).

1 Write the word equation.

potassium carbonate + nitric acid → potassium nitrate + water + carbon dioxide

2 Write the unbalanced equation by replacing the chemical names with the chemical formulas.

$K_2CO_3 + HNO_3 \rightarrow KNO_3 + H_2O + CO_2$

3 Balance each element one by one.

Potassium (K): There are two on the left and only one on the right, so put a 2 in front of the potassium nitrate (KNO_3). This is sensible because you cannot destroy atoms—if you start with 2, you must end up with 2.

$K_2CO_3 + HNO_3 \rightarrow 2KNO_3 + H_2O + CO_2$

Carbon (C): There is one on the left and one on the right, so you don't need to change anything.

Oxygen (O): There are six on the left (three from the K_2CO_3 and three from the HNO_3). However, there are nine on the right:

- 1 from the H_2O
- 2 from the CO_2
- 6 from the $2KNO_3$

Putting a 2 in front of the HNO_3 solves the problem:

$K_2CO_3 + 2HNO_3 \rightarrow 2KNO_3 + H_2O + CO_2$

Now everything balances. (Note that, when trying to balance by adding numbers, this adds multiple lots of everything in the formula. For example, adding a 2 in front of the K_2CO_3 would also balance the oxygen atoms but it would unbalance the potassium and carbon atoms.)

Hydrogen (H): There are now two on the left and two on the right, so this balances.

Nitrogen (N): There are now two on the left and two on the right, so this balances.

4 Double check the numbers of atoms on both sides of the equation.

Reactants = 2 × K, 1 × C, 9 × O, 2 × H, 2 × N

Products = 2 × K, 1 × C, 9 × O, 2 × H, 2 × N

The equation is now balanced.

1 p148 2 p148 5.1

SciFile

In the limelight

Solid calcium oxide (CaO) is also known as lime or quicklime. When it is heated to 2500°C it begins to glow with an extremely bright light. For this reason, it was used in theatres for lighting before the discovery of electricity and is the origin of the term 'in the limelight'.

SCIENCE AS A HUMAN ENDEAVOUR

Nature and development of science

Moving atoms

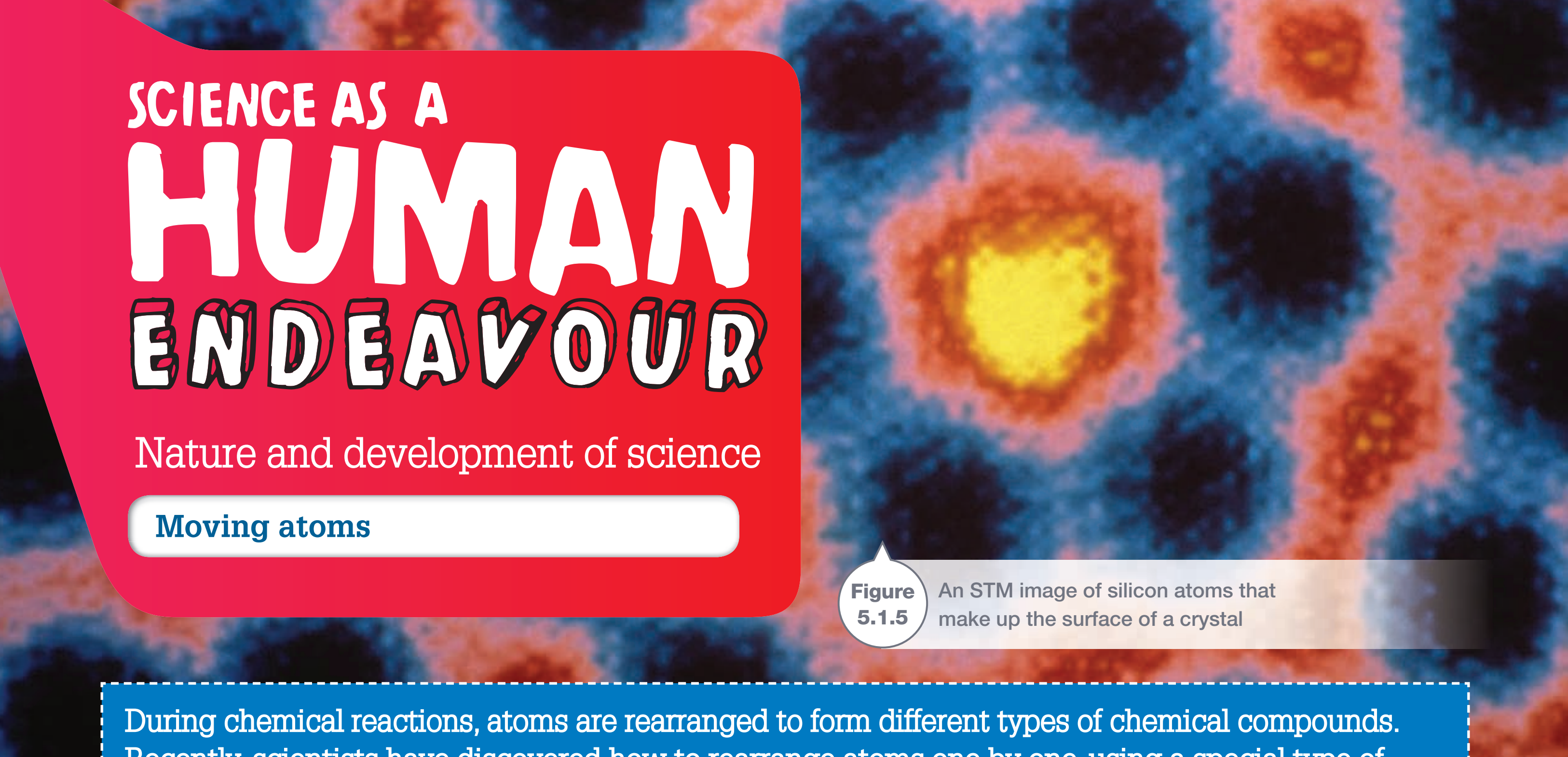


Figure 5.1.5 An STM image of silicon atoms that make up the surface of a crystal

During chemical reactions, atoms are rearranged to form different types of chemical compounds. Recently, scientists have discovered how to rearrange atoms one by one, using a special type of microscope know as a scanning tunnelling microscope (STM).

Scanning tunnelling microscopes

Scanning tunnelling microscopes are the only type of microscope that can create images of atoms like those in Figure 5.1.5. To create an image of atoms, the STM scans the surface of a crystal with a very sharp tip. The sharp tip senses the atoms as bumps on the surface. This is similar to the way in which a sight-impaired person senses the bumps of Braille on a page. In this way, the STM can build up an image of the atoms on the surface line by line, as shown in Figure 5.1.6. Scientists use this technique to get a better understanding of atoms and how they make up crystals.

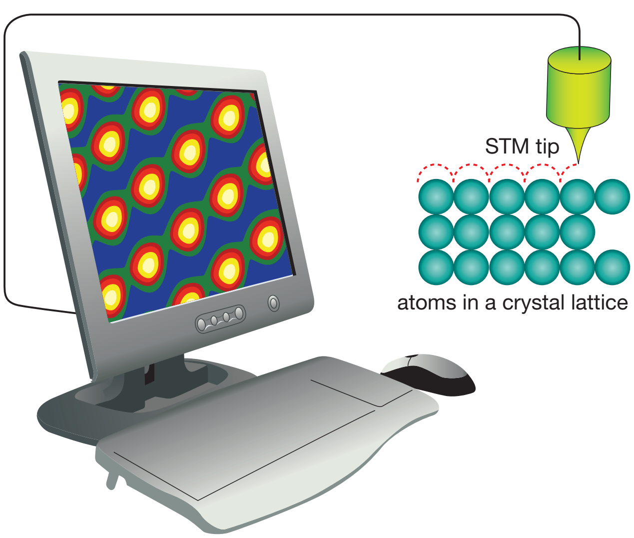


Figure 5.1.6 The STM builds up an image of the surface by sensing the atoms with a very sharp tip.

Moving atoms

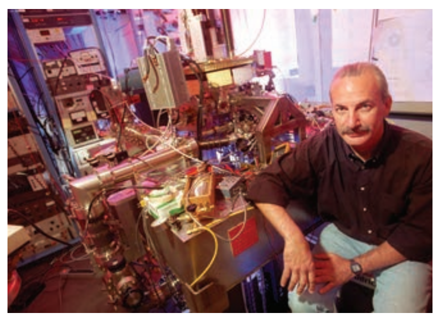


Figure 5.1.7 Professor Donald Eigler was the first person to show that atoms could be moved around with an STM.

In 1989, American nanoscientist Professor Donald Eigler (Figure 5.1.7) showed that the STM could be used to move atoms around one by one. To achieve this, Professor Eigler put iron atoms on a copper surface. He then used the STM tip to push the atoms around the surface and create very precise geometric shapes such as circles, triangles and squares (Figure 5.1.8). Since then, other people have also used the STM to construct designs atom by atom. These include using atoms to write the Japanese characters for the word *atom* as well as the world's smallest stick-figure man.

5.2

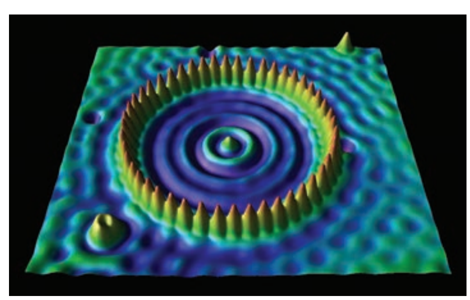


Figure 5.1.8 Using an STM tip, Professor Don Eigler was able to move iron atoms into perfect geometric shapes.

5.1 Unit review

Remembering

1 **State** the general name given to the chemicals that:
- **a** take part in a chemical reaction
- **b** produced by a chemical reaction.

2 **State** what $\rightarrow$ means in chemical equations.

3 **State** the law of conservation of mass.

4 **State** the meaning of the symbols (s), (l), (g) and (aq).

Understanding

5 **Explain** how the law of conservation of mass applies to chemical equations.

6 If a chemical equation contained the formula $2CH_4$, **explain** the meaning of the number 2 and the number 4.

7 **Describe** the difference between NaCl(s) and NaCl(aq).

Applying

8 **a** **Use** word equations to **describe** the following reactions.
b **Identify** the state of the reactants and products.
- **i** When copper is added to dilute nitric acid, copper nitrate, nitrogen monoxide and water are formed.
- **ii** If sulfuric acid is poured onto solid sodium carbonate, bubbles of carbon dioxide are produced, as well as water and sodium sulfate in solution.
- **iii** Magnesium burns easily in oxygen gas, producing magnesium oxide.
- **iv** When sodium metal is placed in water, it reacts to produce sodium hydroxide and hydrogen gas.

9 **Identify** the chemical formula for each of these substances. Include the appropriate state: (aq), (l), (s), (g) at room temperature and pressure.
- **a** carbon dioxide
- **b** dilute sulfuric acid
- **c** hydrogen
- **d** potassium carbonate crystals
- **e** dilute nitric acid
- **f** calcium

10 **Identify** the equation that is *not* balanced.
- **A** $4Al + 3O_2 \rightarrow 2Al_2O_3$
- **B** $Mg + 2HCl \rightarrow MgCl_2 + H_2$
- **C** $2Zn + O_2 \rightarrow 2ZnO$
- **D** $C_5H_{12} + 8O_2 \rightarrow CO_2 + 6H_2O$

Analysing

11 **Calculate** the correct number of reactants and products to balance the following equations.
- **a** $P_4 + O_2 \rightarrow P_2O_5$
- **b** $KClO_3 \rightarrow KCl + O_2$
- **c** $BaO + HNO_3 \rightarrow Ba(NO_3)_2 + H_2O$
- **d** $Pb_3O_4 \rightarrow PbO + O_2$
- **e** $Pb(NO_3)_2 \rightarrow PbO + NO_2 + O_2$

Evaluating

12 Jessica heated some blue copper(II) nitrate crystals in a test-tube. She noticed that brown nitrogen dioxide gas was produced. When a glowing splint was held at the top of the test-tube, it relit, proving that oxygen gas was also produced. A fine black solid, copper(II) oxide, was left in the test-tube.
- **a** **Determine** the reactants and products of the reaction.
- **b** **Deduce** the word equation for this reaction.
- **c** **Deduce** the balanced chemical equation.

13 David added dilute hydrochloric acid to solid calcium carbonate in a beaker. When he weighed the beaker after the bubbling had stopped, he noticed a reduction in mass. **Propose** why his results did not appear to agree with the law of conservation of mass.

Creating

14 Juan burns different masses of magnesium metal (Mg) with oxygen (O_2) to form magnesium oxide (MgO). He measures the mass of the reactants and product before and after, as shown in the table.

Mass of magnesium reacting (g)	Mass of oxygen reacting (g)	Mass of magnesium oxide produced (g)
2.00	0.70	2.70
3.00	1.04	4.04
4.00	1.39	5.39

- **a** **Construct** a word equation for this reaction.
- **b** **Construct** a balanced chemical equation.
- **c** **Modify** the equation to include the states of the reactants and products.
- **d** **Explain** how the above results demonstrate the law of conservation of mass.

Inquiring

1 Investigate the chemical reactions for photosynthesis and respiration. Write the word equations, formula equations and balanced equations for these reactions.

2 Research acid–base reactions. Write the general word equation for acid–base reactions and then three balanced formula equations showing three examples of acid–base reactions.

5.1 Practical activities

1 Reactions with modelling clay

Purpose
To simulate the conservation of mass in chemical reactions using modelling clay models.

Materials
- 3 different colours of modelling clay
- atomic model kit

Procedure
1. Use the atomic model kit to make models of methane (CH_4), oxygen (O_2) and ethane (C_2H_6).
2. Use the modelling clay to create models of the three molecules. Your models should be spheres stuck together to simulate chemical bonds.
3. Without adding, subtracting or splitting any modelling clay atoms, try to simulate how the atoms rearrange according to the unbalanced chemical reaction for burning methane in air:

   methane + oxygen gas → carbon dioxide + water

   $$CH_4 + O_2 \rightarrow CO_2 + H_2O$$
4. Construct one methane molecule and two oxygen molecules and simulate the balanced chemical reaction:

   methane + oxygen gas → carbon dioxide + water

   $$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$$
5. Perform the same simulation for the combustion of ethane (C_2H_6) in oxygen by first simulating the unbalanced equation:

   ethane + oxygen gas → carbon dioxide + water

   $$C_2H_6 + O_2 \rightarrow CO_2 + H_2O$$
6. Simulate the balanced equation with the plasticine models:

   ethane + oxygen gas → carbon dioxide + water

   $$2C_2H_6 + 7O_2 \rightarrow 4CO_2 + 6H_2O$$

Discussion
1. **Describe** what happened when you tried to simulate the unbalanced reactions, listing any excess atoms and those you did not have enough of.
2. **Explain** how a balanced equation gives an accurate description of what is going on during a chemical reaction.

2 Conservation of mass

Purpose
To investigate conservation of mass in a chemical reaction.

SAFETY

Barium nitrate solution is toxic, so avoid contact with eyes, skin and mouth. Wear safety glasses, lab coat and gloves.

Materials
- 20 mL barium nitrate ($BaNO_3$) solution
- 20 mL of sodium sulfate ($NaSO_4$) solution
- 2 × 50 mL beakers
- access to an electronic balance

Procedure
1. Pour approximately 20 mL of barium nitrate into one beaker and approximately 20 mL of sodium sulfate into the other.
2. Measure the weight of each beaker and record your results in the table in the results section.
3. Pour the barium nitrate in beaker 1 into beaker 2.
4. Remeasure the weight of each beaker and record your results.

Results
Copy and complete the following table.

	Beaker 1	Beaker 2	Total
Mass before reaction (g)			
Mass after reaction (g)			

Discussion
1. **Describe** what happened when the two solutions were mixed.
2. The products of this reaction are solid barium sulfate and a sodium nitrate solution. From this information, **construct** a word equation and balanced formula equation for this reaction.
3. **Assess** whether your results agree with the law of conservation of mass. **Explain** your assessment.

5.2 Classifying chemical reactions

There are many different types or classes of chemical reactions. Each type of chemical reaction has something in common. This could be common reactants or products or the way that the chemical reaction takes place. Classifying chemical reactions helps chemists understand how chemical reactions work. By understanding how one chemical reaction works, scientists can understand more about all reactions of the same type.

Decomposition reactions

When a single reactant breaks apart to form several products, the reactant is said to decompose. The general form of a **decomposition reaction** can be written as:

$$XY \rightarrow X + Y$$

An everyday example of a decomposition reaction is the chemical reaction that puts the fizz in soft drinks like the one shown in Figure 5.2.1. Soft drinks contain dissolved carbonic acid (H_2CO_3). When carbonic acid decomposes, it forms water (H_2O) and bubbles of carbon dioxide gas (CO_2). The carbon dioxide gas formed by this reaction remains dissolved in the soft drink until the lid is removed.

The balanced equation for the decomposition of carbonic acid is:

$$\text{carbonic acid} \rightarrow \text{water} + \text{carbon dioxide}$$
$$H_2CO_3(aq) \rightarrow H_2O(l) + CO_2(g)$$

Figure 5.2.1 The decomposition reaction of carbonic acid gives carbonated water its fizz.

INQUIRY science 4 fun

Raisin lava lamp

Can you create a lava lamp with a decomposition reaction?

Yes

Collect this …

- clear fizzy drink such as lemonade or tonic water
- raisins
- clear glass or bottle

Do this …

1 Pour the lemonade into the clear glass or bottle.
2 Add several raisins.

Record this …

Describe what you saw.

Explain why you think this happened.

Thermal decomposition

Some substances will only decompose when heated to high temperatures. This is known as thermal decomposition. Metal carbonates and metal hydrogen carbonates both undergo thermal decomposition.

Thermal decomposition of sodium azide (NaN_3) is a chemical reaction that saves lives every day by inflating vehicle airbags like the one in Figure 5.2.2. When sodium azide is heated, it decomposes into sodium metal and nitrogen gas. The balanced equation is:

$$\text{sodium azide} \rightarrow \text{sodium} + \text{nitrogen gas}$$
$$2NaN_3(s) \rightarrow 2Na(s) + 3N_2(g)$$

Just 100 grams of sodium azide can produce around 56 litres of nitrogen gas in under 0.03 seconds. This reaction rapidly inflates the airbag in the event of a collision.

Figure 5.2.2 The decomposition of sodium azide (NaN_3) saves lives every day.

1 p159

Combination reactions

Combination reactions occur when two reactants combine to form a single product. The general equation for a combination reaction can be written as:

$$X + Y \rightarrow XY$$

Combination reactions are important in industry. For example, a combination reaction is used to create hydrochloric acid for industry and laboratories. First, hydrogen gas (H_2) and chlorine gas (Cl_2) are combined to form hydrogen chloride gas (HCl) in large chemical plants like the one in Figure 5.2.3.

The balanced equation for the combination of hydrogen and chlorine is:

$$\text{hydrogen gas} + \text{chlorine gas} \rightarrow \text{hydrogen chloride gas}$$
$$H_2(g) + Cl_2(g) \rightarrow 2HCl(g)$$

The hydrogen chloride gas that is produced is then bubbled through de-ionised water to produce hydrochloric acid.

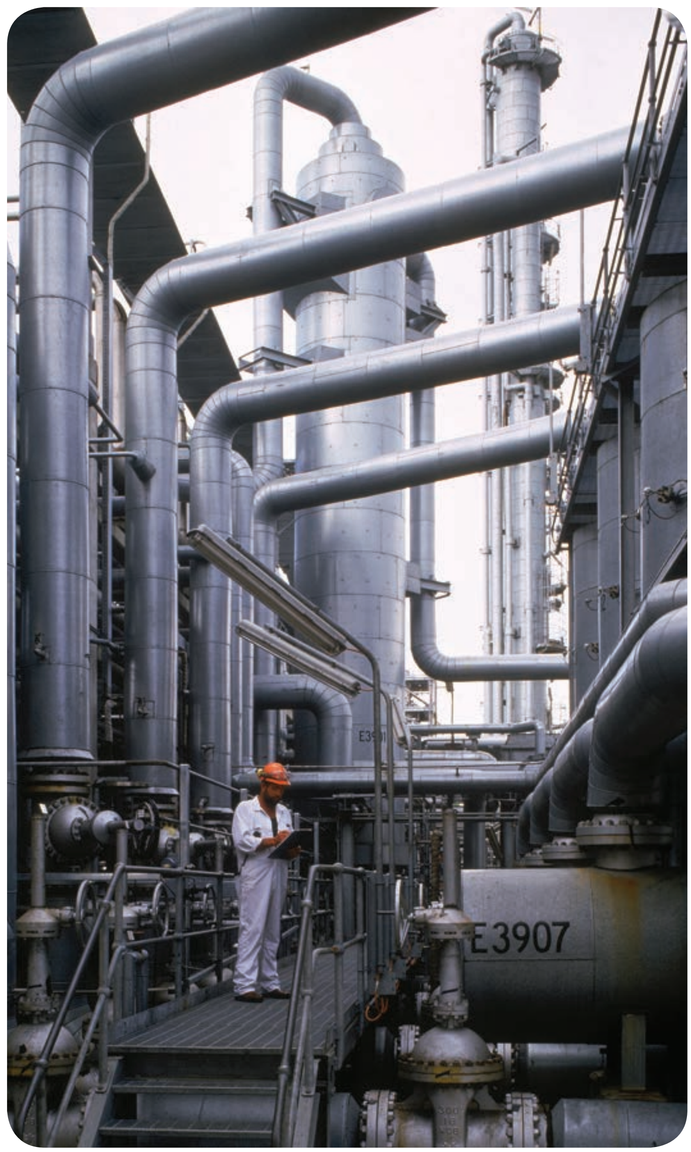


Figure 5.2.3 Hydrochloric acid is manufactured at chemical plants using a combination reaction between hydrogen gas and chlorine gas.

Precipitation reactions

Occasionally when two clear solutions are mixed together, they react to form a solid. The solid is said to precipitate out of the solution. As shown in Figure 5.2.4, most precipitates quickly fall to the bottom of the beaker. These types of reactions are known as **precipitation reactions**. For example, the scale that builds up in kettles, taps and pipes is solid calcium carbonate ($CaCO_3$) that has precipitated out of the tap water.

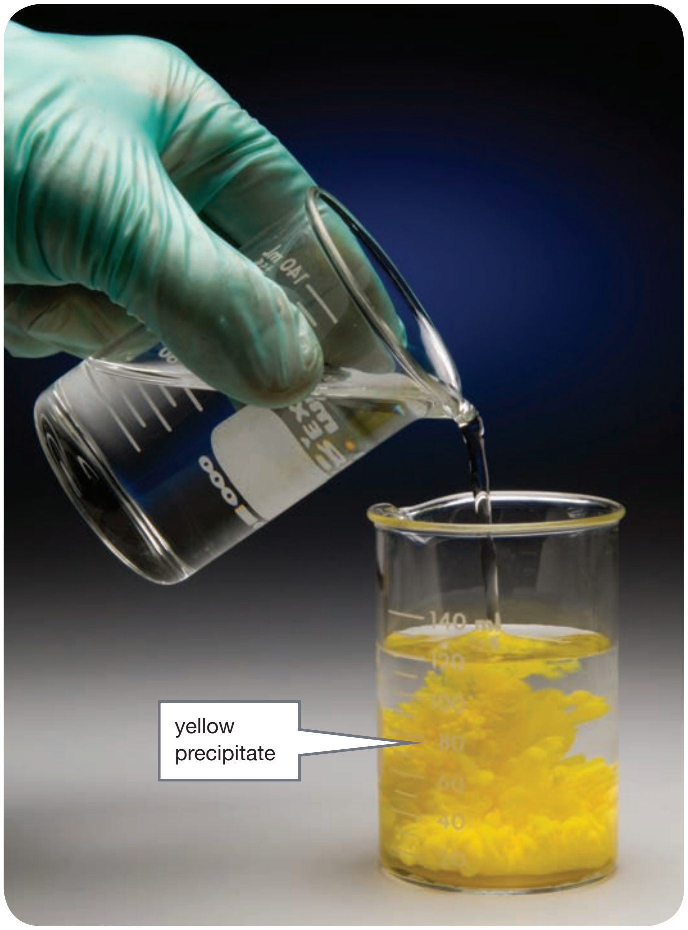
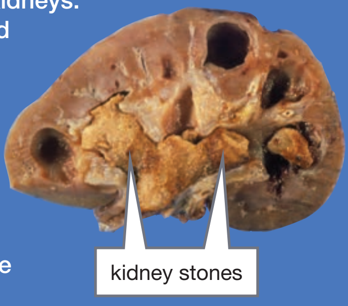


Figure 5.2.4 Precipitation reactions occur when two aqueous solutions react to produce a solid.

Painful precipitates

Your body is full of dissolved compounds. However, sometimes these compounds precipitate out as hard deposits in the kidneys. These deposits, called kidney stones, are extremely painful. Usually kidney stones will pass out of the body with urine. However, in severe cases, the stones may have to be removed surgically or shattered by intense soundwaves.



SciFile

Precipitation reactions and solubility

A precipitation reaction occurs when two **soluble** reactants combine to form an **insoluble** product known as the **precipitate**. A substance is said to be soluble if it dissolves. For example, sugar is soluble in water. A substance is insoluble if it does not dissolve. For example, chalk is insoluble in water.

When a soluble substance is dissolved in water, the particles that make up the substance are spread thinly throughout the solution. The particles are so small and so thinly distributed that they cannot be seen with the naked eye. As a result, the solution appears transparent (clear, not cloudy or murky). It may be coloured like food colouring in water, but you can see through it clearly.

For example, when table salt or sodium chloride (NaCl) dissolves, it breaks apart into positive sodium **ions** (Na^+) and negative chloride ions (Cl^-) as shown in Figure 5.2.5. This is because table salt is an **ionic compound**.

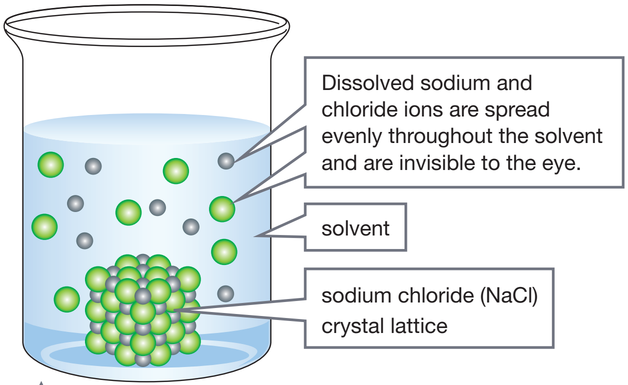


Figure 5.2.5 When sodium chloride dissolves, the lattice breaks apart and the ions distribute through the solution.

Ionic compounds

Ionic compounds are crystalline substances that are made up of a crystal lattice of positive ions (**cations**) and negative ions (**anions**). Ionic compounds are usually hard, brittle and can be brightly coloured like the ones shown in Figure 5.2.6.

Figure 5.2.6 There are many different types of ionic compounds such as salts that come in a wide variety of colours.

The cations that make up the crystal lattice are atoms (or groups of atoms) that have lost electrons and therefore have a positive charge. Anions are atoms (or groups of atoms) that have gained electrons and therefore have a negative charge. Table 5.2.1 lists common cations and anions.

Table 5.2.1 Common cations and anions

		Chemical name	Symbol
Cations	Lost 1 electron	Hydrogen ion	H^+
		Lithium ion	Li^+
		Sodium ion	Na^+
		Potassium ion	K^+
		Ammonium ion	NH_4^+
		Copper(I) ion	Cu^+
	Lost 2 electrons	Calcium ion	Ca^{2+}
		Magnesium ion	Mg^{2+}
		Barium ion	Ba^{2+}
		Copper(II) ion	Cu^{2+}
		Iron(II) ion	Fe^{2+}
	Lost 3 electrons	Iron(III) ion	Fe^{3+}
		Aluminium ion	Al^{3+}
Anions	Gained 1 electron	Fluoride	F^-
		Chloride	Cl^-
		Bromide	Br^-
		Iodide	I^-
		Hydroxide	OH^-
		Nitrate	NO_3^-
		Hydrogen carbonate	HCO_3^-
	Gained 2 electrons	Oxide	O^{2-}
		Sulfide	S^{2-}
		Sulfate	SO_4^{2-}
		Carbonate	CO_3^{2-}
	Gained 3 electrons	Nitride	N^{3-}
		Phosphate	PO_4^{3-}

Naming ionic compounds

The name of an ionic compound is simply the name of the cation together with the name of the anion. For example, barium sulfate ($BaSO_4$) is made up of the barium cation (Ba^{2+}) and the sulfate anion (SO_4^{2-}). In the cases where an atom can have more than one type of ion (such as copper(I), Cu^+, and copper(II), Cu^{2+}), a roman numeral is included in the name of the compound. For example, copper(I) hydroxide ($CuOH$) or copper(II) sulfate ($CuSO_4$).

Ionic compounds are always charge neutral. This is represented in the chemical formula by ensuring that the total charge on the cations balances the total charge on the anions. For example, sodium oxide is made up of sodium ions (Na^+) with a charge of +1 and oxide ions (O^{2-}) with a charge of –2. Therefore, the symbol for sodium oxide is Na_2O. This formula indicates that there needs to be two sodium ions for every oxide ion in the crystal lattice to balance the charge.

Polyatomic ions are ions with more than one atom. Examples are NH_4^+ and SO_4^{2-}. The chemical symbol of these ions is put inside brackets when more than one is needed for a balanced formula. For example, the chemical formula for calcium hydroxide is $Ca(OH)_2$. This indicates that there are two hydroxide ions (OH^-) to balance the charge of each calcium ion (Ca^{2+}).

Predicting precipitation reactions

Scientists use the solubility rules in Table 5.2.2 to predict if a precipitation reaction will occur when two ionic solutions are mixed.

Table 5.2.2 Solubility rules

Negative ions (anions)	Positive ions (cations)	Solubility of compounds
Acetate, CH_3COO^-	All	Soluble
All	Li^+, Na^+, K^+, Rb^+, NH_4^+	Soluble
Chloride, Cl^- Bromide, Br^- Iodide, I^-	Ag^+, Pb^{2+}, Hg^{2+}, Cu^+	Low solubility
	All others	Soluble
Hydroxide, OH^-	Li^+, Na^+, K^+, Rb^+, NH_4^+, Sr^{2+}, Ba^{2+}	Soluble
	All others	Low solubility
Nitrate, NO_3^-	All	Soluble
Phosphate, PO_4^{3-} Carbonate, CO_3^{2-}	Li^+, Na^+, K^+, Rb^+, NH_4^+	Soluble
	All others	Low solubility
Sulfate, SO_4^{2-}	Ca^{2+}, Sr^{2+}, Ba^{2+}, Pb^{2+}	Low solubility
	All others	Soluble
Sulfide, S^{2-}	Li^+, Na^+, K^+, Rb^+, NH_4^+, Be^{2+}, Mg^{2+}, Ca^{2+}, Sr^{2+}, Ba^{2+}	Soluble
	All others	Low solubility

Using the solubility rules, it is possible to predict that a potassium chloride solution and a silver nitrate solution will react to form a silver chloride precipitate and leave behind a potassium nitrate solution (Figure 5.2.7). The balanced equation for this reaction is given by:

$$\text{potassium chloride} + \text{silver nitrate} \rightarrow \text{potassium nitrate} + \text{silver chloride}$$

$$KCl(aq) + AgNO_3(aq) \rightarrow KNO_3(aq) + AgCl(s)$$

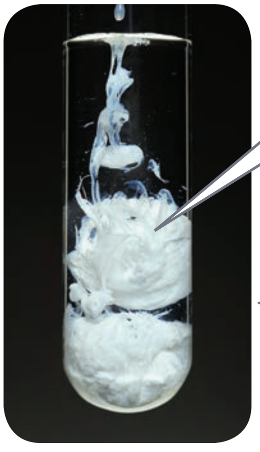


Figure 5.2.7 Silver chloride precipitates out as a white solid when potassium chloride is mixed with silver nitrate.

The solubility rules state that all potassium ionic compounds are soluble and therefore potassium chloride (KCl) is soluble. The rules also state that all nitrates are soluble. Therefore, silver nitrate ($AgNO_3$) must be soluble. Together, these solutions can combine to create potassium nitrate (KNO_3) and silver chloride (AgCl). Potassium nitrate (KNO_3) must be soluble because all potassium ionic compounds and all nitrates are soluble. However, the solubility rules show that silver chloride is only slightly soluble in water and therefore it must precipitate out of the solution as a solid. Using these rules, scientists can predict the outcome of any precipitation reaction.

2 p160

SKILL BUILDER

Predicting precipitation reactions

It is possible to predict the outcome of mixing two solutions by considering the solubility of all the possible combinations of cations and anions. Consider a mixture of solutions of magnesium sulfate ($MgSO_4$) and barium nitrate ($Ba(NO_3)_2$).

Step 1: Swap the cations and anions of the reactants to get the possible products.

Product 1 = Magnesium nitrate ($Mg(NO_3)_2$)

Product 2 = Barium sulfate ($BaSO_4$)

Step 2: Check the solubility of the possible products in Table 5.2.2.

Product 1 = Magnesium nitrate is soluble because all nitrates are soluble.

Product 2 = Barium sulfate is insoluble because all sulfates are soluble except Ba^{2+}, Ca^{2+}, Sr^{2+}, Pb^{2+}.

Step 3: Write the chemical equation for the reaction showing that barium sulfate is a solid precipitate.

$$\text{barium nitrate solution} + \text{magnesium sulfate solution} \rightarrow \text{barium sulfate solid} + \text{magnesium nitrate solution}$$

$$Ba(NO_3)_2(aq) + MgSO_4(aq) \rightarrow BaSO_4(s) + Mg(NO_3)_2(aq)$$

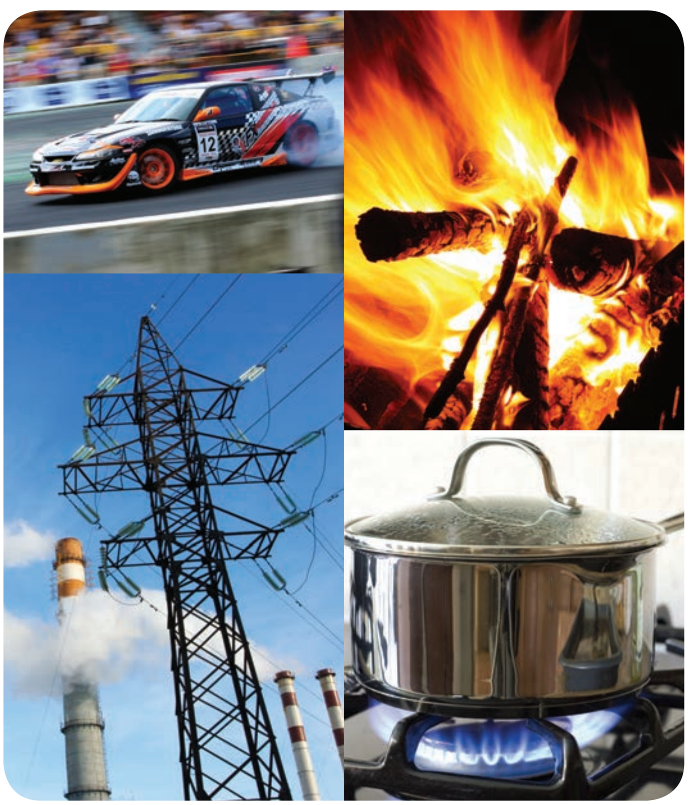


Figure 5.2.8 Combustion is a chemical reaction used every day to cook, provide warmth, run cars and produce electricity.

Oxidation and reduction reactions

Oxidation and reduction reactions include a wide variety of reactions. However, it is not always obvious when an oxidation and reduction reaction has taken place.

Oxidation reactions

If a substance combines with oxygen during a chemical reaction, then you can be sure that an **oxidation** reaction has occurred. Two familiar examples of oxidation reactions are combustion and corrosion.

During **combustion** reactions, substances burn rapidly in oxygen and produce large amounts of heat and light. Lighting a match, burning gas on a stove top, igniting fuel in car engines and burning coal in an electrical power station are all examples of combustion (Figure 5.2.8).

When methane gas is burnt in oxygen on a gas stove or Bunsen burner, the chemical combustion reaction is:

$$\text{methane} + \text{oxygen} \rightarrow \text{carbon dioxide} + \text{water}$$

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$$

In this reaction, the methane is oxidised because it combines with oxygen in the air to form carbon dioxide and water. The oxygen is referred to as the oxidising agent because it has caused the methane to oxidise.

5.3

INQUIRY science 4 fun

Acrobatic flame

Can a flame travel along candle fumes?

Yes

Collect this …

- candle
- box of matches

Do this …

1. Light the candle and let it burn for a couple of minutes.
2. Strike another match and then blow out the candle.
3. Put the flame of the lit match in the fumes coming from the unlit candle.
4. The candle should reignite as if by magic. See how far you can get the flame to jump.

SAFETY

Be very careful when using matches and candles. Be careful to blow matches out and rinse with water to make sure they are out before disposing of them.

Record this …

Describe what you saw.
Explain why you think this happened.

Corrosion occurs when metals react with oxygen to form metal oxides. For example, when iron is exposed to oxygen in the presence of water, it forms rust or iron(III) oxide (Fe_2O_3). You can see it in Figure 5.2.9. The equation for this reaction is:

iron + oxygen gas → iron(III) oxide

$$2Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s)$$

In this case, the iron metal is being oxidised and the oxygen gas is the oxidising agent.

Figure 5.2.9 Rust is the product of an oxidation reaction, more specifically called corrosion.

Reduction reactions

Reduction reactions are the opposite of oxidation reactions. If a compound loses oxygen atoms, you can be sure a reduction reaction has taken place. For example, when copper oxide is heated to very high temperatures in the presence of carbon, the copper(II) oxide loses its oxygen and forms copper metal. The chemical equation for this reaction is:

copper(II) oxide + carbon → copper + carbon dioxide

$$2CuO\ (s) + C\ (s) \rightarrow 2Cu\ (s) + CO_2\ (g)$$

Here, the copper oxide is reduced and carbon is the reducing agent. However, you could also say that the carbon is oxidised and that the copper(II) oxide is the oxidising agent. Oxidation and reduction reactions always occur in pairs. Therefore, they are commonly referred to as **redox** reactions.

3 p161

Redox reactions and electrons

Redox reactions do not always involve oxygen. Redox reactions also include any chemical reaction where electrons are transferred from one substance to another substance. The substance that loses its electrons is said to be oxidised. The substance that gains the electrons is said to be reduced. An easy way to remember this is by the mnemonic OILRIG:

Oxidation **I**s **L**oss **R**eduction **I**s **G**ain

In short:

Oxidation is the gain of oxygen or the loss of electrons.

Reduction is the loss of oxygen or the gain of electrons.

Identifying redox reactions is not always easy. For example, precipitation reactions are not redox reactions. There is no exchange of electrons even though the charged ions in solution come together to form a crystal lattice. Similarly, neutralisation reactions between acids and bases also are not redox reactions because there is no exchange of electrons between atoms.

Metal displacement reactions

Metal displacement reactions are simple redox reactions that involve only the transfer of electrons from one metal atom to another. For example, when zinc metal is placed in a solution of copper sulfate, the zinc atoms form zinc ions (Zn^{2+}) that dissolve into solution. At the same time, the copper(II) ions (Cu^{2+}) form solid copper crystals on the zinc as shown in Figure 5.2.10. Here, the zinc is said to displace the copper from solution.

The chemical equations for this reaction are:

copper(II) ions + zinc → copper + zinc ions

$$Cu^{2+}(aq) + Zn(s) \rightarrow Cu(s) + Zn^{2+}(aq)$$

Figure 5.2.10 When zinc metal is placed in a copper sulfate solution, crystals of pure copper begin to form on the zinc metal strip.

Remembering the mnemonic OILRIG, you can see that the zinc atoms are being oxidised in this reaction because the zinc atoms lose their electrons to the copper(II) ions. Similarly, you can see that the copper(II) ions are reduced in this reaction because they gain electrons from the zinc atoms.

In this redox reaction, electrons have been transferred from the zinc atom to the copper(II) ion. Therefore, the zinc has been oxidised and the copper(II) ion is the oxidising agent. In contrast, the copper(II) ions are reduced by the zinc atoms. So the zinc is the reducing agent.

Activity series

Displacement reactions do not work for every combination of metal and ionic solution. For example, if you put copper metal in a solution of zinc sulfate ($ZnSO_4$), then nothing happens. Whether or not a displacement reaction occurs depends on which atom is better at holding onto its own electrons.

The tendency for metals to react is summarised by the activity series shown in Table 5.2.4.

Table 5.2.4 Activity series for common metals

		Metal	Method of extraction
Lose electrons easily	More reactive	Potassium K	Electrolysis
		Sodium Na	
		Calcium Ca	
		Magnesium Mg	
		Aluminium Al	
		Zinc Zn	Carbon reduction
		Iron Fe	
		Nickel Ni	
		Tin Sn	
		Lead Pb	
		Copper Cu	
		Silver Ag	Native metals
Hold on to electrons more tightly	Less reactive	Gold Au	

Metals closer to the top of the activity series lose their electrons easily. Metals at the bottom of the activity series hold on to their electrons more tightly and can remove electrons from atoms above them. Therefore, a metal displacement reaction will only occur if the solid metal is higher on the activity series table than the metal ions in solution.

According to the activity series, zinc is more likely to lose electrons than copper. Therefore, a copper(II) ion can remove electrons from a zinc atom. But a zinc ion cannot remove electrons from a copper atom.

Extraction of metals

The activity series plays an important role in determining how metals are extracted. Metals that are low on the activity series rarely form compounds. Therefore, metals like gold and silver may exist in their elemental form naturally. Metals in the middle of the activity series tend to form compounds. These compounds can be reduced via a redox reaction to produce the pure metal. Metals at the top end of the activity series have a very high tendency to form compounds or ions. Therefore, the ions of these metals can only be reduced by electrolysis. Electrolysis uses an electrical current to push electrons onto the ions and force them to reduce to the pure element.

Native metals

Pure metals that can be found in nature are known as native metals. Gold is a native metal, as is silver. Therefore, chemical extraction is not required. This is why it is possible to pan for gold or find gold nuggets, as shown in Figure 5.2.11.

Figure 5.2.11 Gold is a native metal, which means it can be found in nature in its elemental form.

However, the fact that gold is not chemically reactive poses a challenge for gold mining companies. These companies must find physical processes for separating the gold from the sand and other minerals. One method is to crush the rocks and sand into a very fine powder. The powder is then put into a tank of water with a chemical similar to washing detergent. When air it bubbled through the liquid, the gold particles stick to the bubbles and float to the surface. The sand and rocks settle to the bottom to be removed as waste. This process of separation is known a **froth flotation**.

Gold and mercury

SciFile

Liquid mercury is one of the few chemicals that can dissolve gold and remove it from crushed ore. The mercury is then heated to evaporate it, leaving the gold behind. Mercury is toxic to humans, so its use in gold mining is banned in many countries, including Australia. However, many poorer countries still use this method, putting workers and people living nearby at risk.

Carbon reduction

Unlike native metals, most other metals exist in nature as compounds. Naturally occurring compounds that are used in the production of pure metals are known as ores. Some common ores are listed in Table 5.2.5.

Table 5.2.5 Common ores and the metals extracted from them

Ore name	Chemical formula	Metal extracted
Zinc blende	ZnS	Zinc Zn
Calamite	$ZnCO_3$	
Haematite	Fe_2O_3	Iron Fe
Magnetite	Fe_3O_4	
Iron pyrites	FeS_2	
Galena	PbS	Lead Pb
Copper pyrites	$CuFeS_2$	Copper Cu
Malachite	$CuCO_3.Cu(OH)_2$	
Cinnabar	HgS	Mercury Hg

The pure metals can usually be extracted from their ores through chemical processes such as carbon reduction.

In **carbon reduction**, the ore is heated to very high temperatures with a source of pure carbon, usually coke. In the blast furnaces used for extracting iron from iron ore (Fe_2O_3), the carbon reduction occurs as a two-step process as shown in Figure 5.2.12.

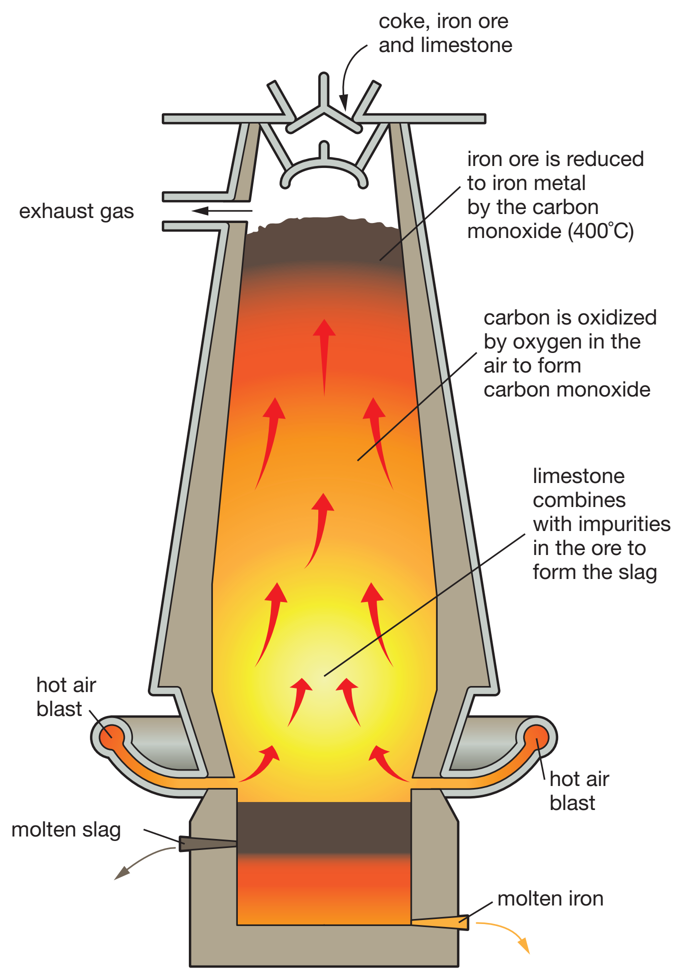


Figure 5.2.12 A blast furnace produces liquid iron by heating iron ore, carbon and limestone to a high temperature.

The iron ore is fed into the top of the blast furnace with coke (pure carbon) and limestone. The carbon is initially oxidised by oxygen in the air to produce carbon monoxide, according to the reaction:

$$\text{carbon} + \text{oxygen gas} \rightarrow \text{carbon monoxide gas}$$
$$2C(s) + O_2(g) \rightarrow 2CO(g)$$

The carbon monoxide then reduces iron ore to produce the iron metal and carbon dioxide.

$$\text{iron(III) oxide} + \text{carbon monoxide gas} \rightarrow \text{iron} + \text{carbon dioxide gas}$$
$$Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(s) + 3CO_2(g)$$

The limestone reacts with the sand and dirt in the ore to form waste known as slag. The slag floats to the top of the molten iron metal and can be drained off separately.

This process of using heat and carbon to extract a metal from its ore is known as **smelting**. Smelting can be used to extract the metals from a wide variety of ores such as zinc blende (ZnS), malachite ($CuCO_3.Cu(OH)_2$), galena (PbS) and tin oxide (SnO).

Figure 5.2.13 Molten iron being poured from a blast furnace

Electrolysis

Carbon reduction is not suitable for the more reactive metals such as sodium, calcium, potassium and magnesium. The most economically viable way of extracting these highly reactive metals is by electrolysis.

Electrolysis is the process in which an electric current is passed through the molten ore (in the liquid phase) in which metal ions are free to move. The electric current pushes electrons onto the metal ions, reducing the ions to atoms of the pure metal. In the mining industry, electrolysis is also called **electrowinning**.

Electrolysis could be used to extract any metal from its ore. However, the process uses huge amounts of electricity and therefore is very expensive. For this reason, electrolysis is used only when cheaper methods (such as carbon reduction) won't work. Common metals extracted by electrolysis are aluminium, sodium, magnesium, calcium and potassium.

Sodium metal is extracted by placing positive and negative electrodes into molten sodium chloride (NaCl) as shown in Figure 5.2.14. Melting the sodium chloride crystal breaks apart the sodium ions and chloride ions, which are then free to move towards the electrodes.

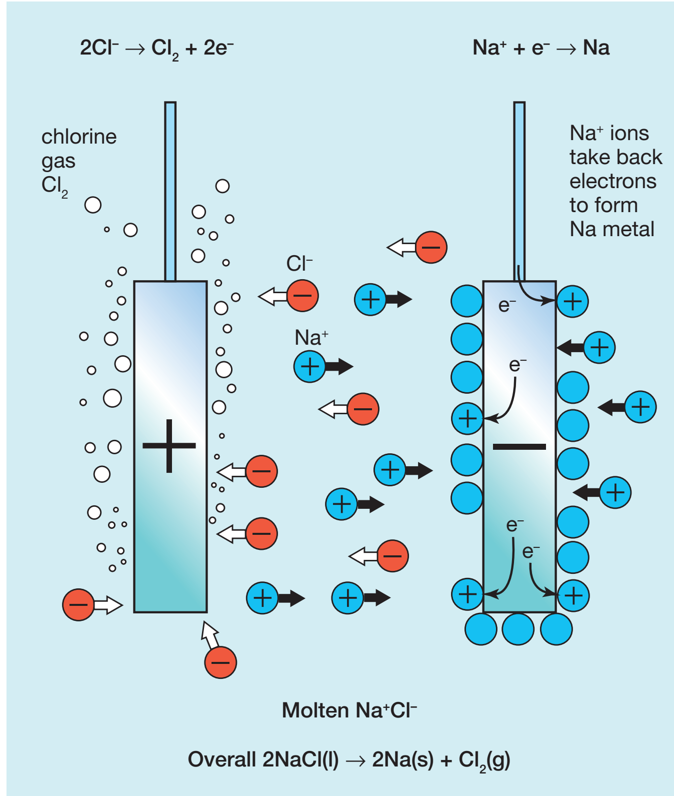


Figure 5.2.14 Pure sodium metal can be extracted from molten sodium chloride through the process of electrolysis.

The positive sodium ions (Na^+) are attracted to the negative electrode, which is charged because of an excess of electrons. When the sodium ions come into contact with the negative electrodes, the excess electrons are pushed onto the sodium ions and the ions are reduced to sodium metal atoms. This can be expressed by the equation:

$$\text{sodium ion} + \text{an electron} \rightarrow \text{sodium}$$
$$Na^+(aq) + e^- \rightarrow Na(s)$$

This is a reduction reaction because the sodium ions have gained electrons. An oxidation reaction occurs at the other electrode. The negative chloride ions are attracted to the positive electrodes, where each chloride ion loses one electron. This causes the chlorine atoms to combine into pairs and form chlorine gas (Cl_2). The gas is chemically stable and bubbles off. This reaction can be expressed by the equation:

$$\text{chloride ion} \rightarrow \text{chlorine gas} + \text{electrons}$$
$$2Cl^-(aq) \rightarrow Cl_2(g) + 2e^-$$

This is an oxidation reaction because the chloride ions have lost electrons. Together, the two equations complete the full redox reaction for the electrolysis of sodium chloride:

$$\text{sodium chloride} \rightarrow \text{sodium} + \text{chlorine gas}$$
$$2NaCl(l) \rightarrow 2Na(s) + Cl_2(g)$$

5.2 Unit review

Remembering

1 **List** two examples each of combination, decomposition, precipitation and redox reactions.

2 **Name** two metals that are:
   a native
   b extracted by carbon reduction
   c extracted by electrolysis.

3 **Name** a positive ion that is non-metallic.

4 **List** names and symbols of three polyatomic ions.

5 **State** whether nitrates are normally soluble or insoluble.

6 **Name** the following ionic compounds.
   a RbBr
   b K_2S
   c BeO
   d Na_3N
   e NH_4Cl
   f LiOH
   g Ag_2CO_3
   h $ZnSO_4$

Understanding

7 **Explain** why oxygen is written as O_2 in chemical reactions rather than just O.

8 **Explain** why combination and decomposition reactions could be considered the reverse or opposite of each other.

9 **Describe** what happens when sodium ions come in contact with the negative electrode in a cell used for electrolysis.

10 **Describe** what observations may be made when a precipitation reaction occurs.

Applying

11 **Identify** a combination reaction that occurs in a blast furnace.

12 **Apply** rules for formula writing to write the formula for calcium chloride.

13 **Identify** if the following reactions are combination, combustion or decomposition reactions.
   a $2KClO_3 \rightarrow 2KCl + 3O_2$
   b $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$
   c $O_2 + 2H_2O \rightarrow 2H_2O_2$

14 **Use** formula equations to rewrite the following word equations.
   a carbon + oxygen → carbon dioxide
   b copper(II) carbonate → copper(II) oxide + carbon dioxide
   c propane (C_3H_8) + oxygen → carbon dioxide + water

15 **Use** word equations to rewrite the following formula equations.
   a $2Mg + O_2 \rightarrow 2MgO$
   b $2H_2O \rightarrow 2H_2 + O_2$
   c $CaO + H_2O \rightarrow Ca(OH)_2$

16 **Use** Table 5.2.2 on page 152 to **predict** the precipitate formed when these solutions are mixed.
   a silver nitrate and sodium chloride
   b mercury(II) nitrate and potassium iodide
   c calcium nitrate and lithium carbonate
   d barium nitrate and sodium sulfate

17 **Use** word and formula equations to **describe** the following redox reactions.
   a aluminium corroding
   b magnesium combusting in oxygen to form magnesium oxide
   c magnesium metal displacing nickel ions from solution

Analysing

18 **Discuss** ways in which combustion is important in our everyday lives.

19 **Calculate** the number of each type of atom in the following chemical formulas.
   a $(NH_4)_2SO_4$
   b $K_2Cr_2O_7$
   c $Ca(OH)_2$

Evaluating

20 **Evaluate** how learning to classify reactions into types has helped you to better understand reactions.

Creating

21 **Construct** a table to **summarise** the different reaction types in this unit. Use your own headings to help you clarify the typical reaction and examples.

Inquiring

1 Explain why the decomposition of mercury oxide (HgO) was a very important reaction for 18th-century chemists Carl Wilhelm Scheele, Joseph Priestley and Antoine-Laurent Lavoisier.

2 Research the ingredients used in gun powder and how the chemical reaction works. Identify which reactants might be oxidizing agents and which might be reducing agents.

3 Research car airbags and find out what heats the sodium azide in a collision.

4 a Design an experiment to determine the contents of three unlabelled beakers, containing a solution of either sodium chloride, barium nitrate or potassium nitrate.
   b Describe what you expect to observe.
   c Write balanced formula equations for each.

5.2 Practical activities

1 Heating metal carbonates

Purpose

To investigate the effect of heat on metal carbonates.

Materials

- 2 large test-tubes
- stopper for test-tube with a delivery tube
- 2 retort stands
- 2 bossheads and clamps
- Bunsen burner and mat
- spatula
- limewater (calcium hydroxide solution $Ca(OH)_2$)
- 5 g copper(II) carbonate

SAFETY

Wear safety glasses. Do not stop heating the copper carbonate for more than about 10 seconds (while the delivery tube is in the limewater) or the limewater will be pushed into the hot test-tube with copper carbonate and it will shatter. One team member must watch this during the heating.

Procedure

1. Use the spatula to put about 5 g (or about 2 cm depth) of copper(II) carbonate ($CuCO_3$) in one of the test-tubes.
2. Fit the test-tube with the stopper and delivery tube.
3. Use the bosshead and clamp to secure the test-tube on an upward angle with the delivery tube pointing down as shown in Figure 5.2.15.
4. Half-fill the second test-tube with limewater and clamp it to the retort stand so that the delivery tube dips into the limewater.
5. Use the Bunsen burner to heat the copper carbonate gently at first and then more strongly.
6. Record your observations but remove the delivery tube from the limewater as soon as the heat is stopped to prevent limewater rising up the delivery tube.

Discussion

1. **Assess** if a chemical reaction has taken place.
2. **Propose** whether the mass of the substance left in the test-tube after heating would be greater or less than the mass of copper carbonate put into the test-tube originally.
3. **Research** how limewater acts as an indicator for carbon dioxide. **Construct** a word equation for the reaction that takes place in the limewater when carbon dioxide is bubbled through it.

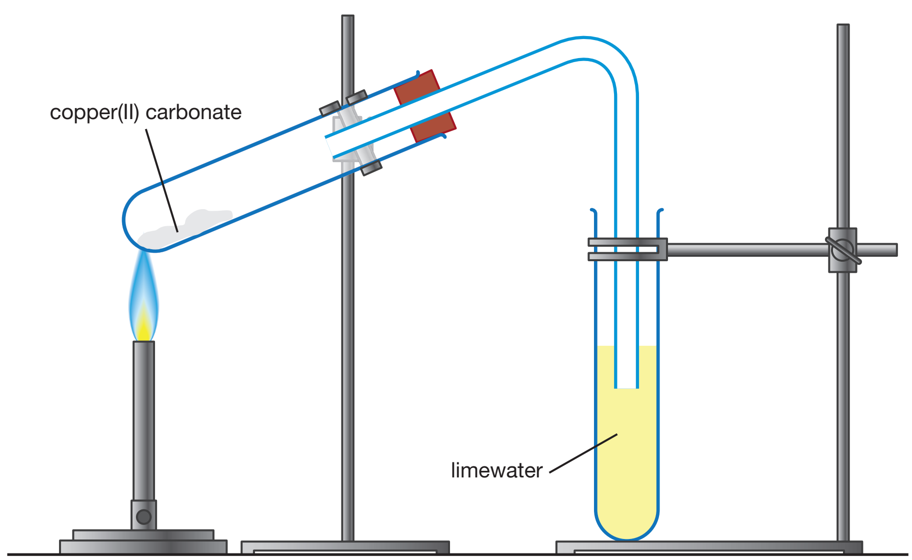


Figure 5.2.15

2 Precipitation reactions

Purpose

To predict and then test particular precipitation reactions.

Material

- Table 5.2.2 (page 152)
- 0.1 M solutions of silver nitrate, sodium carbonate, sodium hydroxide, barium nitrate and copper sulfate (all in 'dropper' bottles)
- 10 small test-tubes and test-tube rack
- gloves

SAFETY

Solutions from this experiment must not be washed down the sink. They should be placed in a clearly marked waste bottle.

Wear gloves and safety glasses at all times.

Procedure

1 Construct a table like the one in the results section. You need 11 rows to record 10 tests.

2 In your table, enter all possible combinations of two of the test solutions.

3 Using Table 5.2.2, predict what should happen in each tube and write your prediction in the table.

4 For your first pair of solutions, place 10 drops of solution A in a test-tube, and then add 10 drops of solution B to the same test-tube. Place the tube in a test-tube rack and record your observations.

5 Repeat step 4 for each pair of solutions.

Results

Copy and complete the table below—you will need 11 rows.

Discussion

1 **Analyse** the solubility rules in Table 5.2.2 on page 152 to work out what has precipitated from each solution.

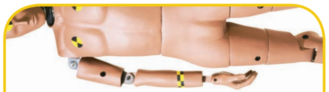
2 **Construct** word equations and formula equations to describe what is happening in each case where a reaction occurred.

Solution A	Solution B	Prediction of precipitation (yes/no)	Observations before mixing	Observations after mixing	Name of precipitate (if any)
Silver nitrate	Sodium carbonate				

3 Carbon reduction

Purpose
To use carbon reduction to extract a metal.

Material
- crucible and lid
- Bunsen burner, tripod and bench mat
- pipe clay triangle
- copper oxide
- powdered carbon
- tin can with top and bottom removed
- 2 teaspoons
- icy-pole stick
- tongs
- stereomicroscope



SAFETY

Be very careful with the hot crucible and pipe clay triangle.

Wear safety glasses at all times.

Procedure
1. Place about one teaspoon of the copper oxide into your crucible.
2. Use a different spoon to add about half a teaspoon of carbon powder. Do not mix up the two spoons.
3. Use the icy-pole stick to thoroughly mix the two together.
4. Use the first spoon (for carbon) to add another teaspoon of carbon to the crucible, but do not mix it. Use the icy-pole stick to spread the carbon into a layer over the top of the mixture in the crucible. Do not mix it with the lower layer.
5. Put a lid on the crucible and stand it in a pipe clay triangle on a tripod stand. Place the tin on the tripod so that it surrounds the crucible. The setup is shown in Figure 5.2.16.
6. Light your Bunsen burner and heat the crucible with a hot flame for about 10 minutes.
7. Let the crucible cool for several minutes. Using the tongs, tip the contents of the crucible onto a tin lid or other container that will not melt. Allow it to cool.

Results
Observe the contents of the crucible under a stereomicroscope or hand lens, and record your observations.

Discussion
1. **Describe** the contents of your crucible before and after heating.
2. **Construct** a word equation and a formula equation to **describe** what happened in the crucible.
3. Carbon reduction using this method works with copper in a laboratory, but not very well with iron. **Propose** a reason for this difference.
4. Another way of using carbon reduction to extract copper from copper oxide is using equipment like that shown in Figure 5.2.17. **Predict** what would happen to the carbon block and the copper oxide powder.

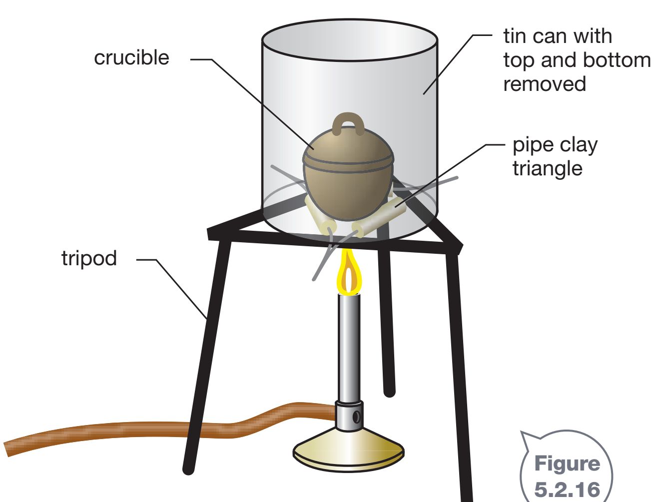


Figure 5.2.16

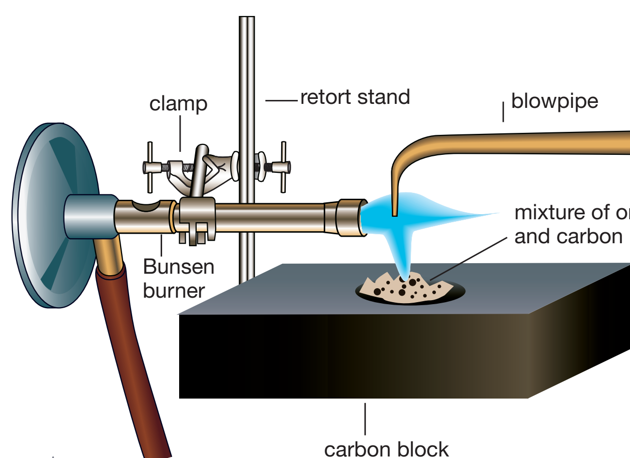


Figure 5.2.17 Alternate way of extracting copper