7 Chemical reactions

Why learn this?

Did you realise that your life is full of chemical reactions? Useful materials such as cloth, paper, ink, ceramics, metals and polymers are all products of chemical reactions. Chemical reactions occur in the cooking and digestion of your food, and they must occur in all the cells of your body to sustain life. Medicines help cure illness by affecting the chemical reactions in your body. Even everyday occurrences — the rusting of an old car, the burning fire of the barbeque and the fizzing of a bath bomb — are the result of chemical compounds interacting.

In this chapter, students will:

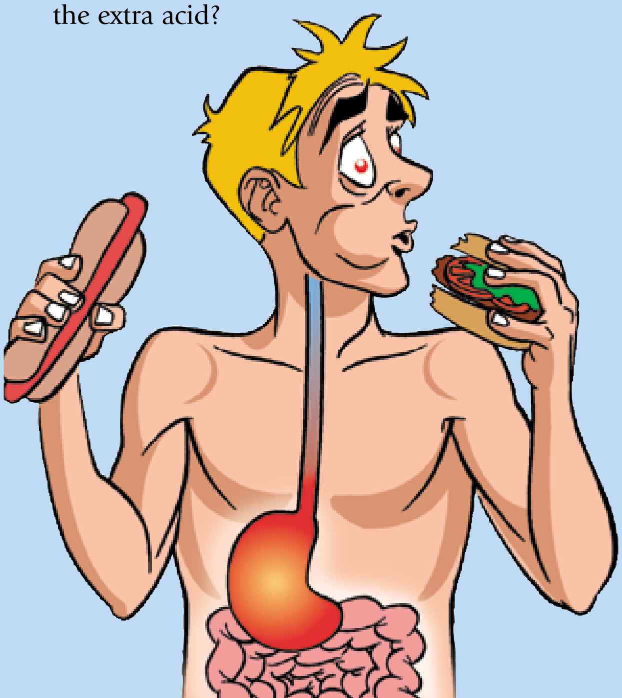
- **7.1** construct word equations to describe chemical reactions
- **7.2** use chemical formulae to write balanced chemical equations
- **7.3** distinguish between exothermic and endothermic reactions
- **7.4** learn the characteristics of acids and bases and the role of indicators, and describe how neutralisation reactions form chemical salts
- **7.5** differentiate between soluble salts and those that will form a precipitate as the result of a chemical reaction
- **7.6** describe how some reactions form precipitates and predict by consideration of the reactants whether precipitates will form
- **7.7** use the activity series to predict the results of displacement reactions
- **7.8** describe a variety of important reactions including corrosion, combustion and decomposition; redox reactions are also examined
- **7.9** explore reactions that release energy in the form of chemiluminescence
- **7.10** appreciate that chemicals may be hazardous and learn how to interpret safety instructions specific to dangerous chemicals.

A glow stick produces light as a result of a chemical reaction between the substances inside the stick and its interior capsule. The energy released in the reaction causes the coloured fluorescent dye to glow.

The chemistry of eating

Preparing, eating and digesting food all involve chemical reactions, many of which you already know about. Answer the following questions to find out what you already know about these important chemical reactions.

1. All of the food that we eat — including meat — begins with the growth of plants.
   (a) What is the name of the chemical reaction that produces the glucose that plants produce?
   (b) Which form of energy is necessary to allow this chemical reaction to take place?
2. The baking of bread makes use of a chemical reaction involving yeast and sugar. The same type of reaction is used in brewing to produce alcohol.
   (a) What is the name of this chemical reaction?
   (b) One of the products of this chemical reaction causes bread to rise while it is being baked. What is the name of this product? (*Clue:* It's a gas.)
3. The chemical digestion of food occurs when chemicals in your body react with the food.
   (a) What name is given to the chemicals that speed up these chemical reactions?
   (b) Much of the food that you eat is broken down into glucose, which takes part in a chemical reaction that occurs in every single cell of your body. What is the name of this chemical reaction, which releases useful energy?
4. Overeating can make your stomach produce too much acid.
   (a) Which type of substance is contained in the products that can be taken to reduce the discomfort and pain caused by the extra acid?
   (b) What is the name of the chemical reaction that provides you with relief from the effects of the extra acid?



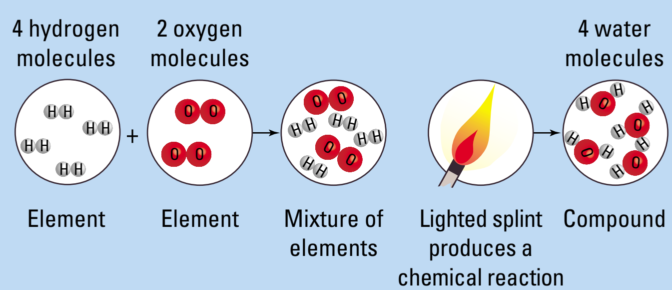
Evidence of chemical reactions

1. Consider the following everyday changes:
- A silver bracelet will tarnish and become black
- A cake rises in an oven as it is cooked
- Butter melts when left out on a counter
- A candle burns
- A bath bomb fizzes when it is added to the water in a bathtub
- Instant coffee dissolves when you add water to it
- Condensed milk heated in a pan until it forms caramel.

   (a) Which of these changes do you think are chemical reactions?
   (b) For each of the changes you selected in (a), what evidence indicates that a chemical reaction has occurred?
   (c) How is a chemical reaction different to a physical change?
   (d) Give three examples of chemical reactions that you encounter in the bathroom.

Inside chemical reactions

Chemical reactions take place when the bonds between atoms are broken and new bonds are formed, creating a new arrangement of atoms and therefore at least one new substance.



1. Explain what happens to the chemical bonds during the chemical reaction between oxygen and hydrogen as illustrated in the diagram above.

Describing chemical reactions

Chemical reactions occur when the bonds between atoms are broken and new bonds are formed creating new combinations of atoms. In fact, once a new substance has been formed it is signalled by observable changes — a change in temperature or colour, the formation of a precipitate or a gas, perhaps even just a different smell being given off. So when you stick some sherbet in your mouth and you feel it fizzing on your tongue, or smell the exhaust fumes from a passing car, you are observing evidence that a chemical reaction has taken place.

Chemistry word equations

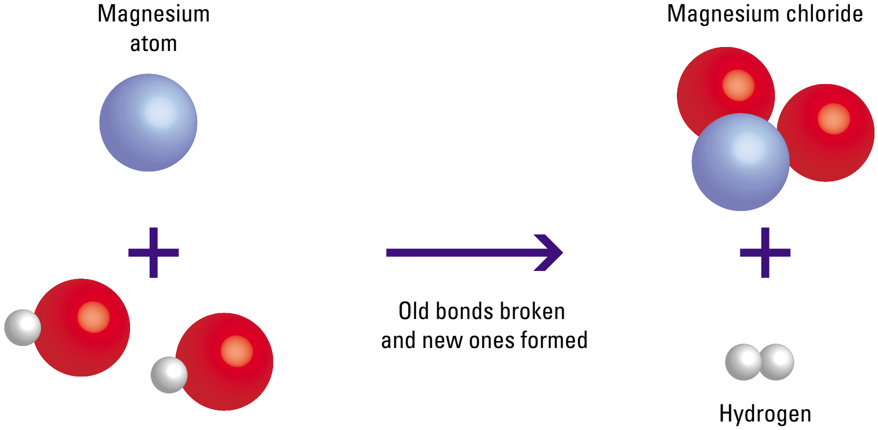
You may recall from your previous studies that chemical reactions can be communicated in the form of a chemical equation. The substances that react together are called the **reactants** while the new substances that are formed are called the **products**.

Chemical equations are written so that the reactants are shown first and are separated by a plus sign (+). The reaction itself is represented by an arrow that points to the products of the reaction (→). The products are also separated by plus signs.

For example, when we add magnesium metal to hydrochloric acid, the new substances magnesium chloride and hydrogen gas are formed. We can represent this in the form of a word equation as shown:

magnesium + hydrochloric acid → magnesium chloride + hydrogen gas

In any chemical reaction, the bonds between the atoms of the reactants have been broken and new bonds formed. In the case of the reaction between magnesium and hydrochloric acid, the ionic bonds between the hydrogen and chlorine atoms which make up hydrochloric acid have been broken; the chlorine atoms have then formed new ionic bonds with the magnesium atoms to form magnesium chloride molecules. At the same time the hydrogen atoms form covalent bonds with the other hydrogen atoms. (*Note:* To brush up on covalent and ionic bonds, go back to chapter 5.)



Notice that the same atoms are present in both the reactants and the products; that is, no new atoms have been introduced during the reaction. The atoms that were present at the beginning of the reaction are still the same ones present at the end. The reaction has simply been a rearrangement of the atoms.

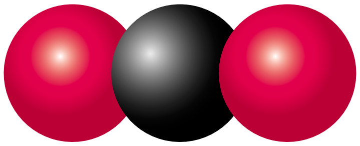
No doubt all of this seems pretty obvious to you, but when the basic mechanics of chemical reactions were first proposed over 200 years ago, this was all very revolutionary.

A burning question

In the eighteenth century, Antoine-Laurent Lavoisier provided the evidence on which these ideas are based. He first considered the way in which a candle seems to disappear as it burns. He did a series of experiments in which he captured all of the gases and soot produced as the candle burned and added their mass to the remaining mass of the candle and melted wax. Lavoisier found that the mass of all of the products produced during the candle's burning was equal to the mass of the original candle. Lavoisier's ideas led to the development of the *Law of Conservation of Mass*, which states that *matter can be neither created nor destroyed during a chemical reaction*.

Lavoisier also provided evidence for the *Law of Constant Proportions*, which states that *a compound, no matter how it is formed, always contains the same relative amounts of each element*.

For example, carbon dioxide (CO_2) always contains the same proportion of carbon and oxygen — about 3 grams of carbon for every 8 grams of oxygen. It does not matter whether the carbon dioxide forms from the reaction of sherbet in your mouth, or from the reaction in the engine of a motorcycle, this proportion is fixed because every molecule of CO_2 is formed when one carbon atom bonds with two oxygen atoms. This law helped to shape our understanding of the way atoms bond together.



A carbon dioxide molecule always contains one carbon atom (black) and two oxygen atoms (red).

Kick-starting reactions

Simply placing two chemicals together does not always mean they will react. For example, hydrogen and oxygen react violently, yet a mixture of these two gases can be stored indefinitely if kept cool in a secure container. Energy must be supplied to start these reactions. Sometimes only a small amount of energy is needed to start (or *initiate*) the reaction. Heat transferred from the surroundings may be enough. Energy may also need to be supplied by an electric current, a beam of light or a Bunsen burner flame. This energy is needed to begin the process of breaking the bonds in the reactants, which allows the atoms to rearrange themselves and form new bonds in the products.

In this case, the word equation is modified to show the word 'heat' written over the reaction arrow:

$$\text{hydrogen} + \text{oxygen} \xrightarrow{\text{heat}} \text{water}$$

INVESTIGATION 7.1

Conserve that mass!

AIM **To carry out a quantitative experiment that demonstrates the Law of Conservation of Mass**

You will need:

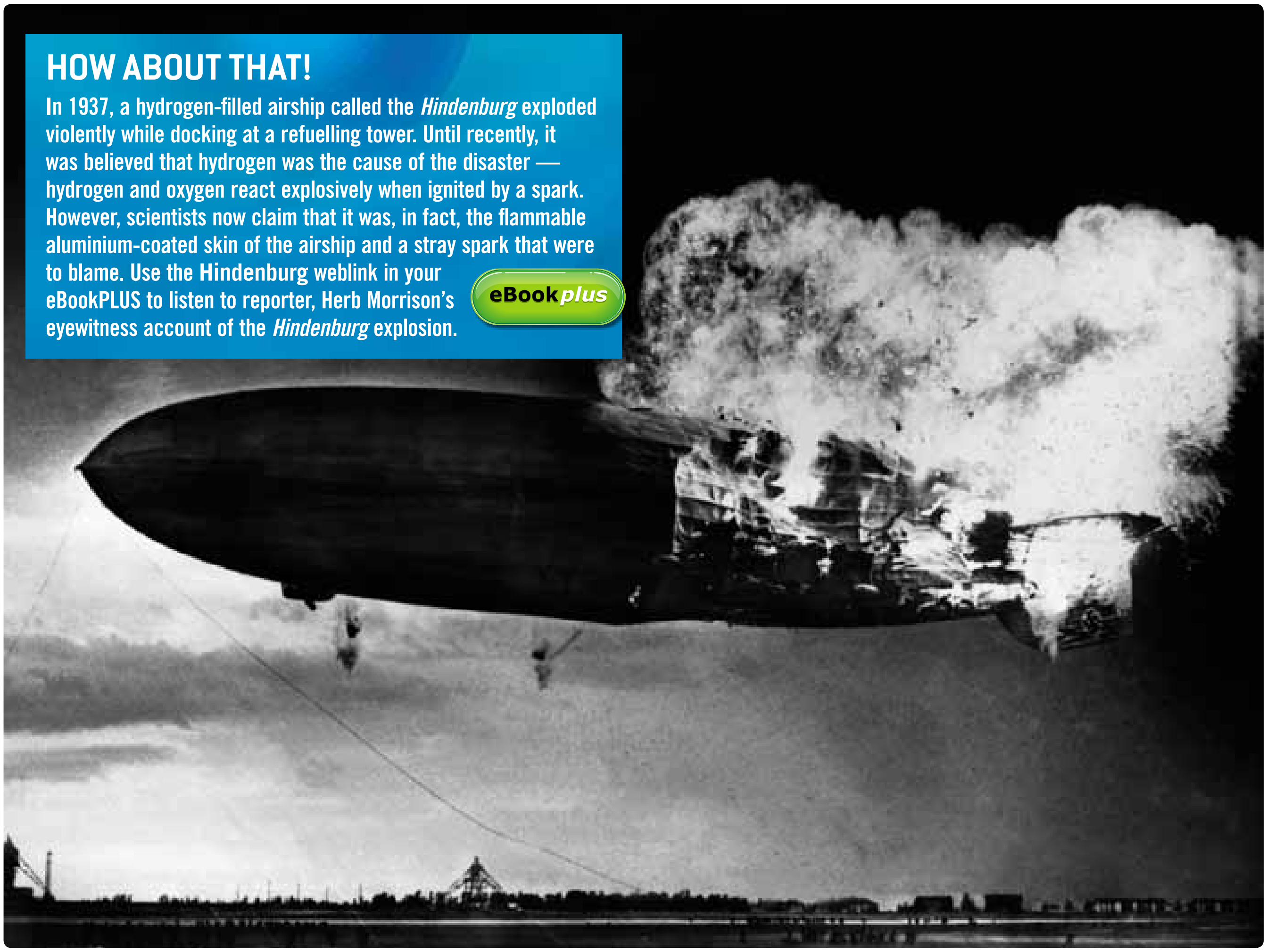
safety glasses
250 mL conical flask
4 Alka-Seltzer tablets
1 balloon
matches
an electronic balance
100 mL measuring cylinder
water

CAUTION *Wear safety glasses.*

- Place the conical flask on the balance and pour in approximately 100 mL of water.
- Place two tablets alongside the conical flask and record the total mass.
- Remove the flask from the balance and drop the tablets into the water. When the reaction is complete, weigh the flask and record the mass.
- Rinse out the flask thoroughly and again add approximately 100 mL of water.
- Place two tablets and a balloon alongside the conical flask and record the total mass.
- Remove the flask from the balance and drop the tablets into the water. Quickly place the balloon on the flask.
- When the reaction is complete, weigh the flask and record the mass. **Do not remove the balloon**.
- After you have recorded the mass, remove the balloon. Light a match and test the gas. Record your observations.

DISCUSSION

1. **Describe what happened during the reaction.**
2. **Which gas do you think filled the balloon?**
3. **Comment on your results of the total mass before and after each reaction. Explain your answer.**
4. **Why do you think it took a long time for the Law of Conservation of Mass to be developed?**



HOW ABOUT THAT!

In 1937, a hydrogen-filled airship called the *Hindenburg* exploded violently while docking at a refuelling tower. Until recently, it was believed that hydrogen was the cause of the disaster — hydrogen and oxygen react explosively when ignited by a spark. However, scientists now claim that it was, in fact, the flammable aluminium-coated skin of the airship and a stray spark that were to blame. Use the Hindenburg weblink in your eBookPLUS to listen to reporter, Herb Morrison's eyewitness account of the *Hindenburg* explosion.

ACTIVITIES

REMEMBER

1 **Recall** the name for chemicals that: (a) combine in a reaction; (b) are formed in a chemical reaction.

2 **Describe** what happens to the atoms in substances that take part in chemical reactions.

3 State the Law of Conservation of Mass and **explain** in your own words what it means.

4 **Recall** the Law of Constant Proportions.

5 Energy can be required to start a reaction. **Recall** three possible sources of this energy.

THINK

6 A piece of paper is weighed on an accurate balance, and then burnt, leaving a pile of ashes. The ashes are collected and weighed on the same balance. Would you expect the mass of the ashes to be the same as the mass of the paper before it was burnt? **Explain** your answer in terms of the products produced.

7 **Explain** how, when a piece of steel wool burns, the mass of the blackened material is greater than the original mass of the steel wool.

8 When a piece of sodium metal is added to a trough of cold water, a rapid and energetic reaction occurs.
(a) **Explain** where the energy to initiate the reaction comes from.
(b) Is energy absorbed or released in this reaction?

9 A chemical reaction is described by the following word equation:

sodium sulfate + **barium chloride** → **barium sulfate** + ______

Identify the name of the second product.

10 The picture at right shows what happens when you add lead nitrate (which is colourless) to potassium iodide (which is also colourless). **Explain** where the yellow substance came from.

INVESTIGATE

11 Find out more about Antoine-Laurent Lavoisier, his work and why he lost his head during the French Revolution.

7.2

The language of chemical equations

In order to communicate, people need to speak a common language. Chemists also use a common language to communicate with each other — a chemical language made up of chemical equations instead of sentences and chemical formulae in place of words!

In chapter 5, we learned how to write chemical formulae for ionic and covalent compounds. These chemical formulae indicate the numbers of atoms of each element present in a particular compound.

The chemical formulae of compounds obey the Law of Constant Proportions. That is, every compound has a fixed relative number of each type of atom. For example, all pure water (H_2O) has two hydrogen atoms for every oxygen atom. Sodium chloride (NaCl) has one sodium atom for each chlorine atom. The formula for copper sulfate is $CuSO_4$ because for every copper atom there is one sulfur atom and four oxygen atoms.

So far, we have just used word equations to describe chemical reactions. Now, we will move on to using the appropriate chemical formulae for the reactants and the products in a chemical equation.

Writing chemical equations

Writing equations involves some simple mathematics and a knowledge of chemical formulae. Chemical equations are set out in the same way as word equations, with the reactants to the left of the arrow and products to the right of the arrow. However, they are different from word equations in three ways:

1. formulae are used to represent the chemicals involved
2. the physical states of the chemicals are often included
3. numbers are written in front of the formulae in order to balance the numbers of atoms on each side of the equation.

There are a few rules to observe in this game of balancing equations. They are described below with a worked example of the explosive reaction between hydrogen gas and oxygen gas. Make sure you read through the rules very carefully before you play the game.

Game rules

GAME RULE 1. Know your products

The products of a reaction must be known from either **observation** or reliable sources (such as chemists) to tell us the products. For example, it is well known that the product of the reaction between hydrogen gas and oxygen gas is water vapour (gas).

GAME RULE 2. Know your formulae

You need to know the formulae of all the reactants and products. For example:

- formula of hydrogen gas H_2
- formula of oxygen gas O_2
- formula of water vapour H_2O.

Remember! Because each substance has only one correct chemical formula, it cannot be changed by altering the subscript numbers.

GAME RULE 3. Write down the formulae

The formulae must be written according to the word equation, with reactants on the left-hand side of the arrow and products on the right-hand side.

$$H_2 + O_2 \rightarrow H_2O$$

GAME RULE 4. Balance the numbers of atoms

First, make a list of the elements present in the formulae under the heading 'Element', as shown below. Then count up how many atoms are represented by the formula of each element under the headings 'Reactants' and 'Products'.

Element	Reactants	Products
H	2	2
O	2	1

You can see that there are not enough oxygen atoms on the product side of the equation. The only way this can be adjusted is by writing numbers in front of the chemical formulae. When we write a number **in front** of a formula, it **multiplies all the atoms** in that formula. Let's increase the number of oxygen atoms on the product side by placing a 2 in front of the formula for water.

$$H_2 + O_2 \rightarrow 2H_2O$$

Recounting the atoms we find:

Element	Reactants	Products
H	2	4
O	2	2

The oxygen atoms are now balanced, but the hydrogen atoms are not. Let's try writing a 2 in front of hydrogen's formula on the reactant side to increase the number of hydrogen atoms.

$$2H_2 + O_2 \rightarrow 2H_2O$$

Counting the atoms again we find:

Element	Reactants	Products
H	4	4
O	2	2

The numbers of each of the elements are the same on both sides of the equation. The equation is balanced!

GAME RULE 5. Include the states

To indicate the physical state of each chemical involved in the reaction, the following symbols are used.

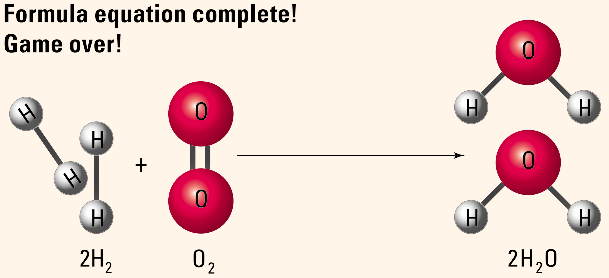
- Solid (s)
- Liquid (l)
- Gas (g)

The symbol (aq) is used to represent an **aqueous solution** of a substance. An aqueous solution is obtained when a substance is dissolved in water.

Write the correct symbol representing the physical state of each reactant and product:

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

Formulae correct!
Number of atoms balanced!
States correct!
Formula equation complete!
Game over!



The reaction between hydrogen and oxygen

Play the game

- Write a word equation and an equation using formulae for each of the six reactions listed. An example is provided on the next page. See the tables at right for the correct formulae.

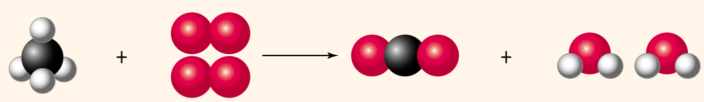
1. Carbon monoxide gas and oxygen gas react to form carbon dioxide gas.
2. Sodium hydroxide solution and hydrochloric acid solution react to form sodium chloride solution and water.
3. Mercury metal and oxygen gas react to form solid mercury(II) oxide.
4. Magnesium metal and hydrochloric acid solution react to form hydrogen gas and magnesium chloride solution.
5. Sodium metal and water react to form hydrogen gas and sodium hydroxide solution.
6. Copper sulfate solution and sodium hydroxide solution react to form solid copper hydroxide and sodium sulfate solution.

The formulae of some common ionic compounds

Compound	Formula
Sodium hydroxide	NaOH
Sodium chloride	NaCl
Magnesium chloride	$MgCl_2$
Copper hydroxide	$Cu(OH)_2$
Sodium sulfate	Na_2SO_4
Copper sulfate	$CuSO_4$
Sodium hydrogen carbonate	$NaHCO_3$
Mercury(II) oxide	HgO
Sodium citrate	$C_6H_5O_7Na_3$

The formulae of some common covalent substances

Compound	Formula
Water	H_2O
Citric acid	$C_6H_8O_7$
Carbon dioxide	CO_2
Oxygen	O_2
Hydrochloric acid	HCl
Carbon monoxide	CO
Hydrogen	H_2

<table>
<tr><th>Balancing a chemical equation</th><th>Example (Methane gas will burn in air. This is an example of a combustion reaction. This type of reaction produces CO_2 and H_2O.)</th></tr>
<tr><td>Step 1: Start with the word equation and name all of the reactants and products.</td><td>Methane gas + oxygen gas → carbon dioxide + water</td></tr>
<tr><td>Step 2: Replace the words in the word equation with formulae and rewrite the equation.</td><td>Methane gas = CH_4 Oxygen gas = O_2 (reactants)
Carbon dioxide = CO_2 Water vapour = H_2O (products)
$\mathbf{CH_4 + O_2 \rightarrow CO_2 + H_2O}$</td></tr>
<tr><td>Step 3: Count the number of atoms of each element (represented by the formulae of the reactants and products).</td><td><table>
<tr><th>Element</th><th>Reactants</th><th>Products</th></tr>
<tr><td>C</td><td>1</td><td>1</td></tr>
<tr><td>H</td><td>4</td><td>2</td></tr>
<tr><td>O</td><td>2</td><td>3</td></tr>
</table></td></tr>
<tr><td>Step 4: If the number of atoms of each element is the same on both sides of the equation, the equation is already balanced. If not, numbers need to be placed in front of one or more of the formulae to balance the equation. These numbers are called coefficients and they multiply all of the atoms in the formula.</td><td>To balance the hydrogen atoms, put a 2 in front of H_2O:
$\mathbf{CH_4 + O_2 \rightarrow CO_2 + 2H_2O}$.
The oxygen atoms can be balanced by putting a 2 in front of the O_2 on the left:
$\mathbf{CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O}$.
The equation is now balanced. It can be checked by counting the number of atoms of each element on both sides of the new equation.<table>
<tr><th>Element</th><th>Reactants</th><th>Products</th></tr>
<tr><td>C</td><td>1</td><td>1</td></tr>
<tr><td>H</td><td>4</td><td>4</td></tr>
<tr><td>O</td><td>4</td><td>4</td></tr>
</table></td></tr>
<tr><td>Step 5: Add physical state symbols.</td><td>$\mathbf{CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)}$
</td></tr>
</table>

ACTIVITIES

REMEMBER

1 **Describe** three differences between word equations and equations in which formulae are used.

2 **Recall** how different states (solid, liquid and gas) are indicated in a chemical equation.

3 **Define** the term 'aqueous solution' and describe how it is represented in a chemical equation.

4 Write balanced chemical equations for the following reactions:
(a) hydrochloric acid reacts with copper to form copper chloride and hydrogen
(b) water and sodium nitrate are formed when nitric acid reacts with sodium hydroxide.

THINK

5 **Identify** what symbols you would use in a chemical equation to represent the metals: iron, mercury, zinc and aluminium (both chemical symbol and state).

6 Write a balanced equation using formulae for the reaction that occurs when you eat a sherbet lolly. These sweets commonly contain citric acid and sodium hydrogen carbonate. In the mouth, these chemicals dissolve in your saliva and then react together to form sodium citrate solution, carbon dioxide gas and water.

7 **Explain** why it is necessary to balance chemical equations.

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8 Test your ability to balance chemical equations by completing the **Checking for balance** interactivity in your eBookPLUS. **int-0677**

9 To learn more about balancing chemical equations use the **Balancing equations** weblink in your eBookPLUS.

work sheets
7.1 Chemical equations
7.2 Balancing chemical equations

7.3

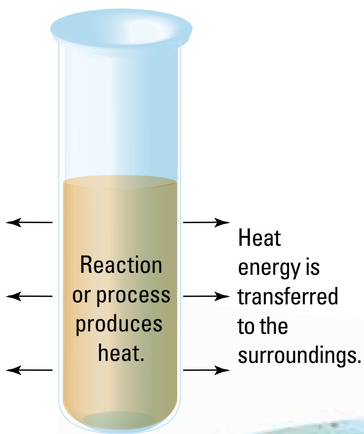
Hot and cold changes

When chemical reactions occur, bonds are broken between the atoms of the reactants and new bonds are formed when the atoms reform into the product molecules. The breaking and reforming of bonds involves energy being released or absorbed. Often, when a chemical reaction is taking place, you can feel the effect of these energy changes as a change in temperature of the chemicals.

Bringing on the heat

A reaction or process that gives out heat is said to be **exothermic.** The term exothermic comes from the Greek words *exo* meaning 'outside' and *therme* meaning 'heat'. An exothermic process may be a chemical reaction in which new products are formed, the dissolving of a substance, the crystallisation of a substance or the change in state a substance undergoes when it turns from a gas into a liquid or from a liquid into a gas.

The energy produced is called **thermal** energy. The thermal energy produced in an exothermic reaction comes from the chemical energy stored in the bonds between the atoms of the reactants. When these bonds break, the energy is released.



Exothermic processes are used in a number of different heat packs that can be used for warming when there is no easy access to electricity. Some of these processes involve chemical reactions while others are physical changes. These heat packs may be used to keep hands warm in a cold environment or to warm muscles before athletics or physiotherapy in case of injury.

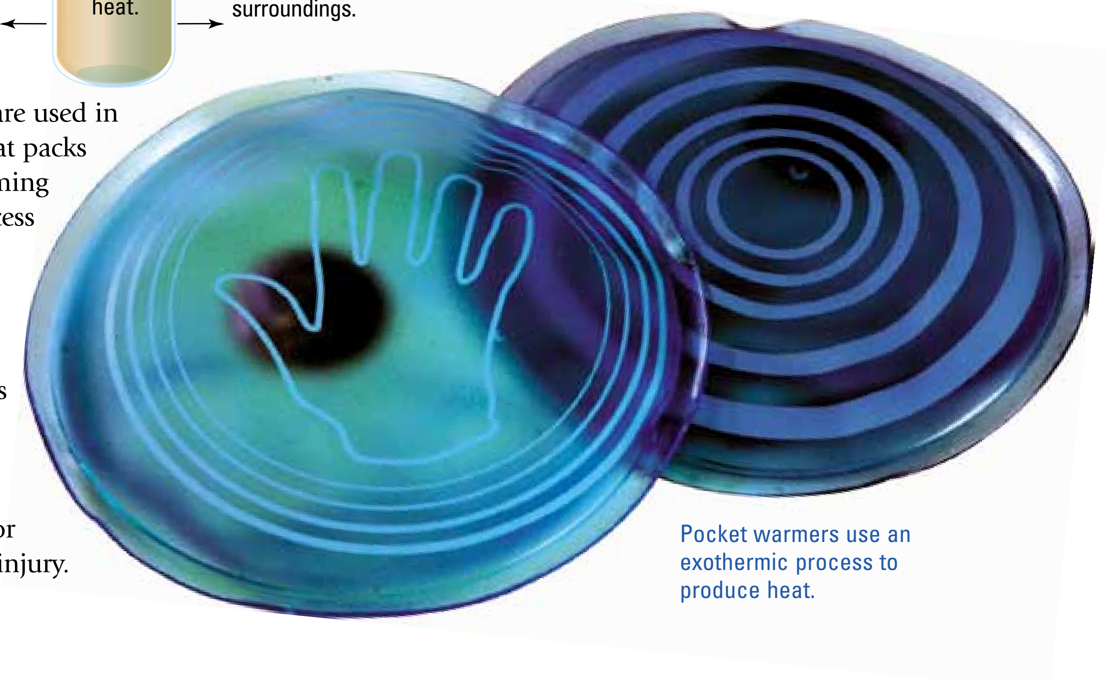
One type of heat pack contains iron powder, sawdust, salt, charcoal and water. When the pack is shaken, the iron combines with the oxygen in the air to form iron oxide:

$$4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s)$$

The sawdust absorbs and spreads the heat in the same way that wheat does in heat bags that are heated in the microwave. The energy produced by this chemical reaction can keep the pack warm for up to 20 hours.

Some exothermic processes are not chemical reactions because no new substance has been formed. Instead, heat is released by the breaking or formation of bonds caused by the rearrangement of the compound molecules into new forms or states of matter.

Most of the pocket warmers that are sold in camping stores are filled with sodium acetate and have a small metal disc inside. When the metal disk in the pocket warmer is cracked, small crystals of sodium acetate are released into the sodium acetate solution in the bag. The solution in the bag is **supersaturated** — that means that it has more solute dissolved in it than normal. When the crystals enter the solution, it triggers **crystallisation** in the solution, releasing thermal energy. This causes the pocket warmer to heat and keep your pockets — and the hands in them — toasty warm for a few hours. This is described as an exothermic process rather than an exothermic reaction as no new substances are formed.



Pocket warmers use an exothermic process to produce heat.

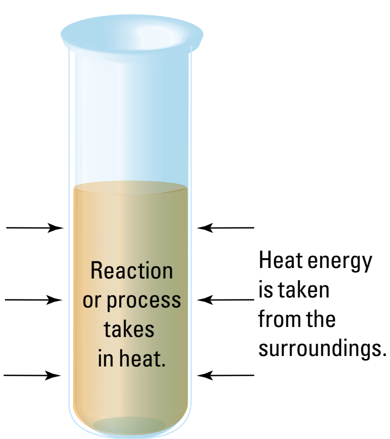
Yet another type of heat pack contains calcium chloride powder in a bag with an inner bag of water. When the bag of water bursts, heat is released as the bonds between the atoms in the calcium chloride break, forming ions of calcium and chlorine which then spread throughout the water, forming an aqueous solution of calcium chloride:

$$CaCl_2(s) \xrightarrow{\text{water}} CaCl_2(aq)$$

Cooling it!

Endothermic reactions are the opposite of exothermic reactions. These reactions take in thermal energy from their environment in order to make or break the chemical bonds between atoms. Because they remove heat from their environment, endothermic reactions reduce the temperature of the reacting chemicals.

As with exothermic processes, endothermic processes can be chemical reactions, or physical changes such as when a substance is dissolved or when there is a change of state. Some reactions can get so cold that water vapour from the air freezes on the outside of the container. This type of reaction can come in very handy at times.



INVESTIGATION 7.2

Exothermic processes

AIM **To investigate some reactions that produce heat**

You will need:

3 large test tubes and test-tube rack
10 mL measuring cylinder
balance
spatula and watch glass
thermometer (–10 °C to 110 °C)
stirring rod
magnesium ribbon
sandpaper
0.5 mol L^{-1} hydrochloric acid
0.5 mol L^{-1} sodium hydroxide
anhydrous copper(II) sulfate

Construct a table, like the one below, to record the temperature changes in each of the following experiments.

CAUTION *Wear safety glasses.*

Part 1: Magnesium in hydrochloric acid

- Put 10 mL of 0.5 mol L^{-1} hydrochloric acid in a test tube and place it in the test-tube rack. Place a thermometer in the test tube and record the temperature.
- Clean a 5 cm piece of magnesium ribbon with the sandpaper until it is shiny on both sides. Coil the magnesium ribbon and put it in the test tube of hydrochloric acid.
- Record the final temperature of the solution as the magnesium reacts with the hydrochloric acid.

Part 2: Hydrochloric acid and sodium hydroxide

- Put 10 mL of 0.5 mol L^{-1} hydrochloric acid in a test tube and place it in the test-tube rack.
- Place a thermometer in the test tube and record the temperature.
- Add 10 mL of 0.5 mol L^{-1} sodium hydroxide solution to the test tube and record the final temperature of the solution as the hydrochloric acid reacts with the sodium hydroxide.

Part 3: Dissolving anhydrous copper(II) sulfate in water

- Put 10 mL of water in a test tube and place it in the test-tube rack. Record the temperature of the water.
- Use a balance to weigh out 2 g of anhydrous copper(II) sulfate. Add this to the test tube and stir gently.
- Record the final temperature of the solution as the copper(II) sulfate dissolves.

DISCUSSION

1. **Complete the table by calculating the increase of temperature in each case.**
2. **Sometimes the energy given out is described as the heat of reaction and in other cases it is called the heat of solution. For each of the above, decide in which category they belong.**

Exothermic processes

Chemical process	Initial temperature (°C)	Final temperature (°C)	Increase in temperature (°C)

INVESTIGATION 7.3

Endothermic processes

AIM To investigate some reactions that absorb heat from their surroundings

You will need:

3 large test tubes and a test-tube rack
10 mL measuring cylinder
balance
spatula and watch glass
thermometer (–10 °C to 110 °C)
stirring rod
potassium nitrate
sodium thiosulfate.

For class demonstration:

barium hydroxide octahydrate
ammonium thiocyanate
2 × 250 mL beakers
stirring rod
electronic thermometer
wooden block
wash bottle.

Construct a table like the one below in which to record the temperature changes in each of the following experiments.

CAUTION *Wear safety glasses.*

Part 3 is a class demonstration using a fume hood. Ammonia is one of the products.

Endothermic processes

Chemical process	Initial temperature (°C)	Final temperature (°C)	Decrease in temperature (°C)

Part 1: Dissolving potassium nitrate in water

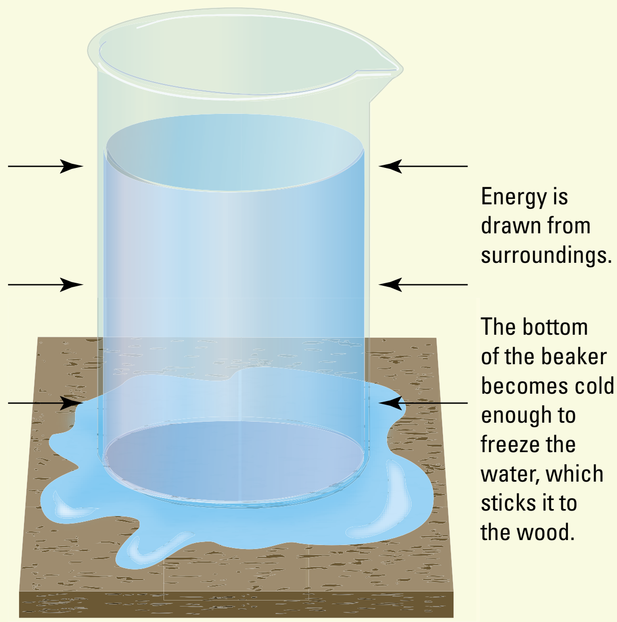
- Put 10 mL of water in a test tube and place it in the test-tube rack. Record the temperature of the water.
- Use a balance to weigh out 2 g of potassium nitrate. Add this to the test tube and stir gently. Record the final temperature of the solution as the potassium nitrate dissolves.

Part 2: Dissolving sodium thiosulfate in water

- Put 10 mL of water in a test tube and place it in the test-tube rack. Record the temperature of the water.
- Use a balance to weigh out 2 g of sodium thiosulfate. Add this to the test tube and stir gently.
- Record the final temperature of the solution as the sodium thiosulfate dissolves.

Part 3: Class demonstration — barium hydroxide octahydrate and ammonium thiocyanate

- Use a balance to weigh out 32 g of barium hydroxide octahydrate into a beaker and record its temperature. Place in the fume hood.
- Using a second beaker, weigh out 17 g of ammonium thiocyanate.
- In the fume hood, mix the ammonium thiocyanate into the first beaker and, using a stirring rod, blend the two solids.
- Squirt some water on the wooden block and sit the beaker on the block.
- Record the final temperature of the mixture with an electronic thermometer.



Ammonium thiocyanate is mixed with barium hydroxide.

DISCUSSION

1. **Complete the table by calculating the decrease of temperature in each case.**
2. **In which of the above were new substances produced?**
3. **Which of the above were (a) chemical reactions (b) physical changes?**

Ice on tap

Athletes use the endothermic processes in instant ice packs to cool injured or swollen muscles. One type of icepack has powdered ammonium nitrate in a bag with an inner bag of water. When the bag of water is burst, the ammonium nitrate dissolves in the water and then draws energy from its surroundings, making the icepack cold. When held against the body, the icepack draws heat from the injured area. This process is described by the equation:

$$NH_4NO_3(s) \xrightarrow{\text{water}} NH_4NO_3(aq)$$

Remember that, as no new substances have been formed, this is not actually a chemical reaction.

ACTIVITIES

REMEMBER

1 How are exothermic reactions different from endothermic reactions?

2 In a chemical reaction in which energy is absorbed from the surroundings, where does the extra energy go?

3 **Explain** why the chemical process that takes place in an icepack containing ammonium nitrate is not a chemical reaction.

4 **Explain** what the term supersaturated means.

THINK

5 The exothermic reaction produced in pocket warmers can be reversed by placing the bag in hot water. What do you think you would see happening in the bag during this time?

6 **Explain** why the combustion (burning) of methane is an exothermic reaction.

7 Are the chemical reactions described below exothermic or endothermic?
(a) Dilute hydrochloric acid is added to dilute sodium hydroxide in a test tube. They react to produce sodium chloride and water. After the reaction, the test tube feels very warm.
(b) As garden compost decomposes, the compost heap gets warmer.
(c) Barium hydroxide and ammonium thiosulfate solutions are mixed and the temperature drops enough to freeze water.

8 Instant hot compresses are used by athletes to warm torn muscles. They relieve pain and speed up the healing process. Some of these hot compresses contain calcium chloride powder and an inner bag of water. When the inner bag bursts, the calcium chloride dissolves in the water and releases energy.
(a) Is the chemical process that takes place in the compress endothermic or exothermic?
(b) How does the energy stored in the chemical bonds of the product compare with the energy stored in the chemical bonds of the calcium chloride and water?
(c) Write an equation to describe this process.

9 Are explosions endothermic or exothermic reactions? **Explain** your answer.

10 In exothermic chemical reactions, energy is released. Why is energy not included in the chemical equations that describe the reactions?

INVESTIGATE

11 Alfred Nobel, the Swedish chemist, made the explosive nitroglycerine much safer.
(a) Use the internet to find out how he did this.
(b) Why did he have to move his laboratory?
(c) What did Alfred Nobel do with his fortune? Why?

12 Use a yearbook or the internet to find out who won the most recent Nobel prizes for Chemistry, Physics and Medicine. Write a short biography about one of the laureates. (The winners of Nobel prizes are referred to as laureates. The Nobel prizes are announced in October of each year.)

Acids and bases

Chemical reactions involving acids and bases play an important role in our lives. They occur in the kitchen, in the laundry, in the garden, in swimming pools and even inside the body. This unit revises your knowledge about acids, bases and neutralisation reactions.

Acids

Acids are corrosive substances. That means that they react with solid substances including metals, marble and even the enamel of your teeth, effectively 'eating' them away. When strong acids such as the sulfuric acid used in car batteries come into contact with the skin, they break down the proteins and fats in the living tissues. This reaction is exothermic so the area heats up dramatically. At best, the result is an acid burn that is slow to heal.

However, not all acids are strong acids. The weaker acids in ant and bee stings may cause pain but there is little lasting damage to the tissues around the sting site. Others, like the acids in citrus fruits and vinegar, are safe — even pleasant — to taste. Acids such as these are routinely added to food to give it the distinctive sour taste associated with acids — in fact, the word 'acid' comes from the Latin word *acidus* meaning sour.

Acids can also be used to preserve food or to react with other substances in food to produce carbon dioxide gas which makes foods like sponge cake light and fluffy.

Bases

Bases have a bitter taste and feel slippery or soapy to touch. Some bases are very corrosive, especially caustic soda which will break down fat, hair and vegetable matter. Other bases are used in shampoos, toothpaste, and cleaning agents like dishwashing liquid and cloudy ammonia. Bases that can be dissolved in water are called alkalis.

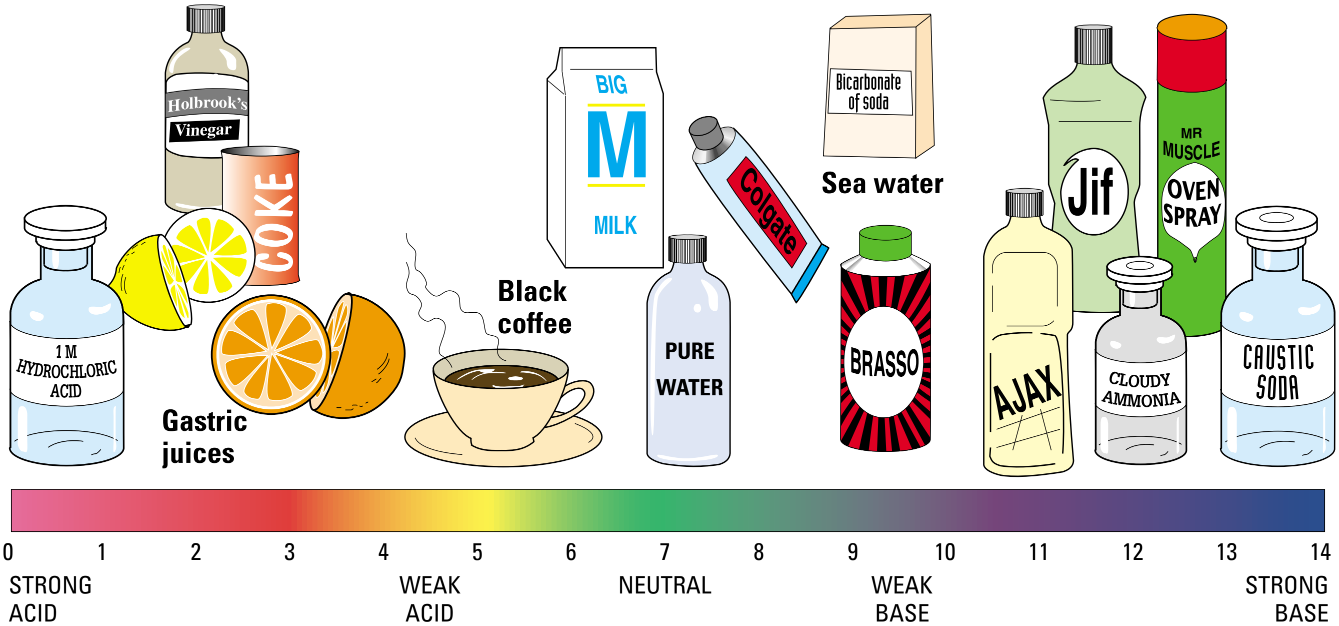
Describing acids and bases

You can describe how acidic or basic a substance is by using the numbers on the pH scale. The pH scale ranges from 0 to 14. Low pH numbers (less than pH 7) mean that substances are acidic. High pH numbers (more than pH 7) mean that substances are basic. If a substance has a pH of 7 it is said to be **neutral** — neither acidic nor basic. This is shown on the pH scale at the top of the opposite page. Acids and bases can be graded from strong to weak. For example, a strong acid has a very low pH (pH 0 or 1) and a strong base has a very high pH (pH 13 or 14).

Common acids and bases

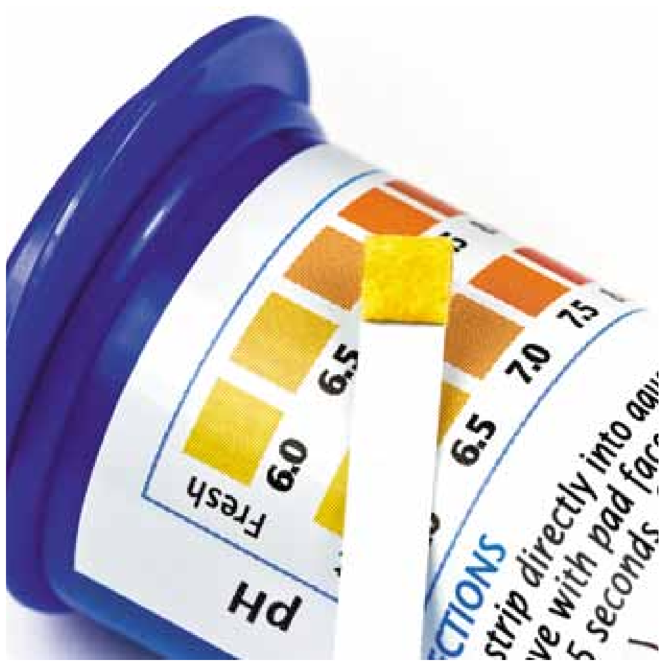
Acid	Uses
Hydrochloric acid	• To clean the surface of iron during its manufacture • Food processing • The manufacture of other chemicals • Oil recovery
Nitric acid	• The manufacture of fertilisers, dyes, drugs and explosives
Sulfuric acid	• The manufacture of fertilisers, plastics, paints, drugs, detergents and paper • Petroleum refining and metallurgy
Citric acid	• Present in citrus fruits such as oranges and lemons • In the food industry and in the manufacture of some pharmaceuticals
Carbonic acid	• Formed when carbon dioxide gas dissolves in water: present in fizzy drinks
Acetic acid	• Found in vinegar • In the production of other chemicals, including aspirin

Base	Uses
Sodium hydroxide (caustic soda)	• In the manufacture of soap • As a cleaning agent
Ammonia	• In the manufacture of fertilisers and in cleaning agents
Sodium bicarbonate	• To make cakes rise when they cook



The pH values of some common substances

A pH meter measures the number of hydrogen ions that are in the liquid and uses this to calculate the pH which is displayed on the screen.

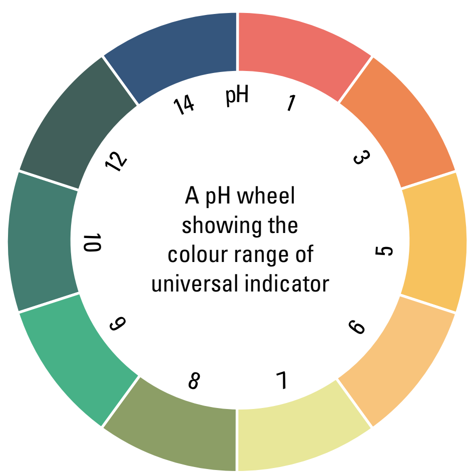


Different indicators respond to different ranges of pH.

Measuring pH

The pH of a substance can be measured using a pH meter or liquids called **indicators.** Indicators change colour depending upon the pH of the substance in which they are placed. The colour ranges of a number of different indicators over the pH range 1–10 are shown below.

Universal indicator is a mixture of indicator substances that changes colour over the full range of pH from 1 to 14 as the strength of an acid or base changes. It is pink in strong acid (pH 1), blue in strong base (pH 14) and green in neutral solutions (pH 7).



The colour range of universal indicator. It is pink in strong acid (pH 1), blue in strong base (pH 14) and green in neutral solutions (pH 7).

Neutralisation

To neutralise something means to stop it from having an effect. Acids and bases are able to counteract or neutralise each other. To stop acid from having an effect, a base can be added to it. For example, the pain caused by the acidic sting of an ant can be relieved by washing the area with an alkali solution containing a weak base such as sodium bicarbonate (baking soda). Similarly, to stop a base from having an effect, an acid can be added. This is why the pain caused by the alkali in the sting of a wasp can be relieved by pouring a weak acid such as vinegar over the area.

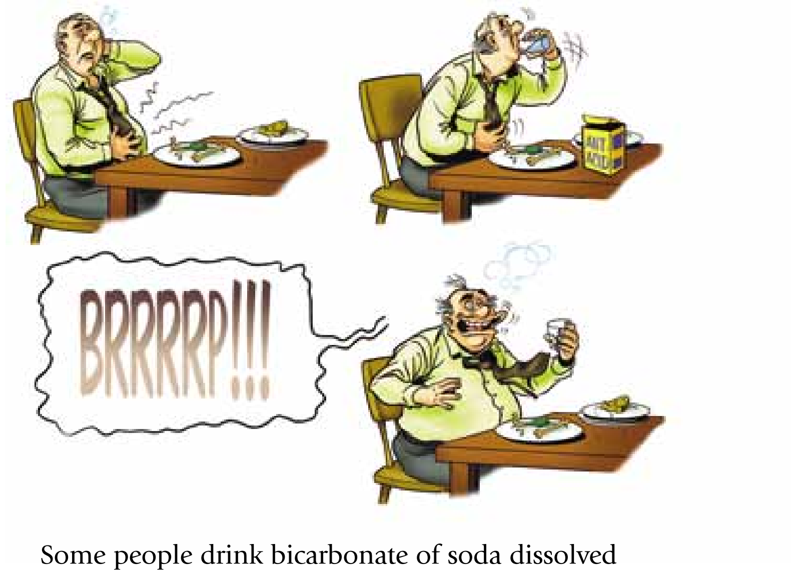
When an acid and a base react with each other, the products are water and a salt (and sometimes a gas) which are neutral (pH7). As a result, such a reaction is called a neutralisation reaction. When hydrochloric acid reacts with sodium hydroxide, water and the salt sodium chloride are produced:

$$NaOH(aq) + HCl(aq) \rightarrow H_2O(l) + NaCl(aq)$$

Neutralisation in action

Indigestion

The hydrochloric acid in your stomach helps to break down the food you eat. It is a very strong acid, with a pH of less than 1.5. But if you eat too quickly, or eat too much of the wrong food, the contents of your stomach become even more acidic. You feel a burning sensation because of the corrosive properties of the acid.



Some people drink bicarbonate of soda dissolved in water to neutralise the acid. However, this produces carbon dioxide gas which builds up in the stomach, making these people feel very uncomfortable and having to burp a lot to relieve the pressure of the gas.

$$HCl(aq) + NaHCO_3(aq) \rightarrow NaCl(aq) + H_2O(l) + CO_2(g)$$

Hydrochloric acid + sodium hydrogen carbonate → sodium chloride + water + carbon dioxide

The most effective way of relieving the indigestion is to take antacid tablets. The active ingredients in modern antacid tablets are weak bases such as aluminium hydroxide or magnesium hydroxide, which neutralise the acid without producing the carbon dioxide gas along with it!

INVESTIGATION 7.4

Antacids in action

AIM To use an antacid to neutralise dilute hydrochloric acid and monitor the reaction with a pH indicator

You will need:

Petri dish
electronic balance
spatula
antacid powder
0.1 moL L^{-1} hydrochloric acid
250 mL conical flask
100 mL measuring cylinder
methyl orange indicator
white tile or white paper

- Measure and record the mass of the Petri dish.
- Add a small amount of antacid powder to the dish and record the mass of the antacid and Petri dish.
- Calculate the mass of the powder.
- Add 50 mL of the dilute hydrochloric acid to the 250 mL flask.
- Add 3 drops of methyl orange indicator.
- Place the flask mixture on the white tile (or paper) and use the spatula to slowly add antacid from the Petri dish bit by bit. Swirl the flask to mix. Stop adding antacid when the colour changes from red to orange.
- Measure and record the mass of the Petri dish and its contents (the unused antacid).

DISCUSSION

1. **What was the mass of the antacid powder?**
2. **What colour change occurs when the methyl orange indicator is in the acid?**
3. **By subtraction, calculate the mass of antacid used to neutralise 50 mL of dilute hydrochloric acid.**
4. **How does your result agree with other groups in your class? Suggest reasons for the similarities or differences between your results.**
5. **Use your results to calculate how much antacid you would need to neutralise 500 mL of dilute hydrochloric acid.**

$$2HCl(aq) + Mg(OH)_2(aq) \rightarrow MgCl_2(aq) + 2H_2O(l)$$

Hydrochloric acid + magnesium hydroxide → magnesium chloride + water

$$3HCl(aq) + Al(OH)_3(aq) \rightarrow AlCl_3(aq) + 3H_2O(l)$$

Hydrochloric acid + aluminium hydroxide → aluminium chloride + water

Of course, care should always be taken when using antacid tablets. If you are using these too often, the stomach reacts by making more acid because it needs the acid to digest proteins. If you then suddenly stop taking the antacid tablets, the increased amount of acid can cause serious health problems.

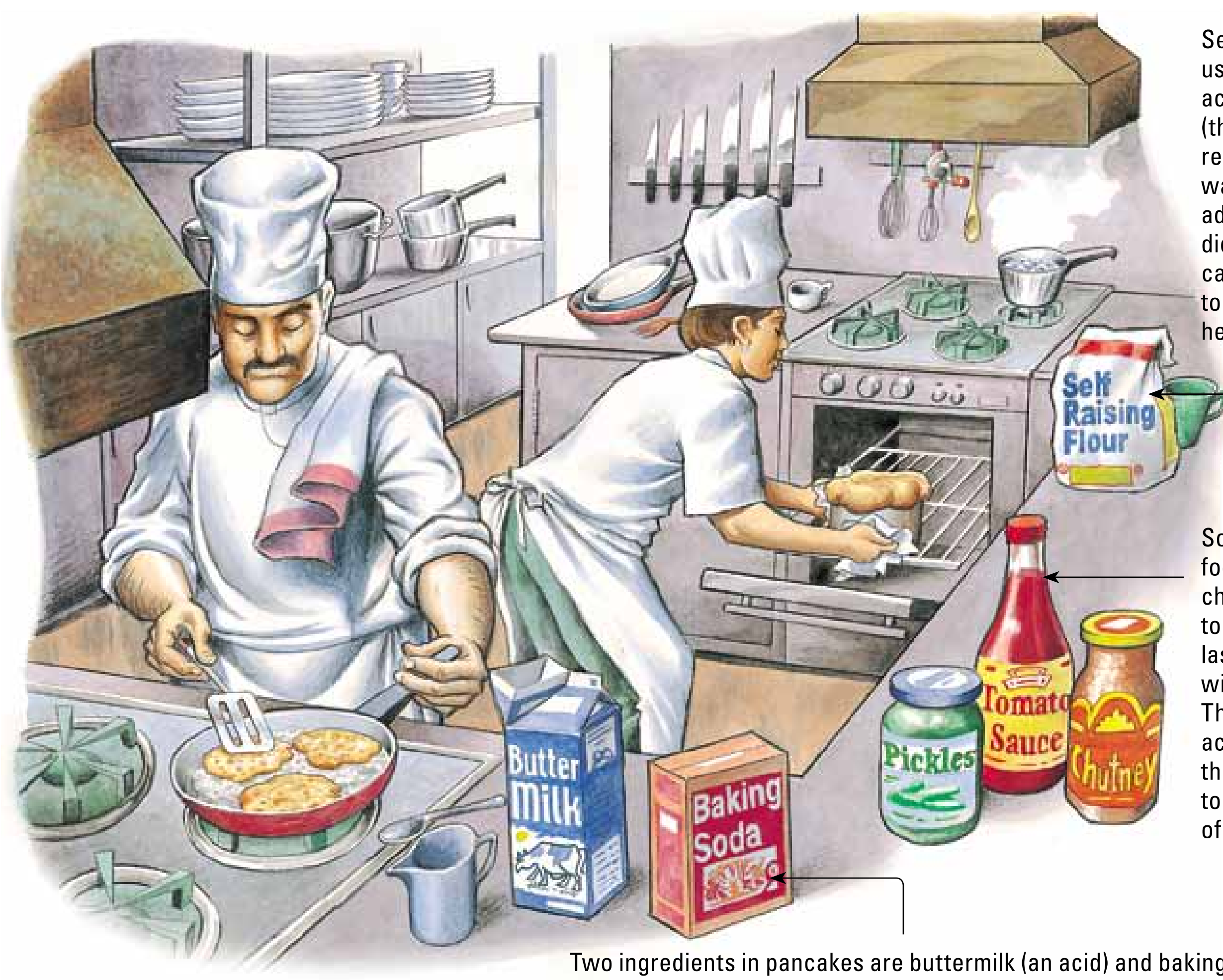
How does your garden grow?

When the soil is too acidic, it can be 'sweetened' (made less acidic) by mixing an alkaline substance called **garden lime** into it. Garden limes are available in a variety of forms but usually contain some combination of calcium hydroxide, calcium carbonate or magnesium carbonate that neutralise some of the soil's acidity. Many gardeners prefer the limes that contain magnesium carbonate as plants use magnesium to make the green chlorophyll in their leaves that allows them to perform photosynthesis.

> **HOW ABOUT THAT!**
> Indigestion is a very old problem and, with not an antacid tablet in sight, many ancient civilisations developed some very interesting ways of treating it. The Egyptians recommended crushing a hog's tooth and putting it inside four sugar cakes which you would then eat. They also believed that swallowing crushed and powdered limestone would help. Interestingly enough, scientists believe that this last one may have worked because the main component of limestone is calcium carbonate. Favourite treatments in medieval England were a bit yummier and involved chewing mint leaves or drinking teas brewed from thyme or pomegranate peel.

What's cooking?

Many acids and bases are found in the kitchen. They help to preserve food and make foods rise when baking. Acids and bases are also found in cleaners used around the house.



Self-raising flour usually contains an acid and baking soda (the base) which react together when water or milk is added. The carbon dioxide produced causes the mixture to rise when it is heated.

Some bottled foods, like pickles, chutney and tomato sauce last a long time without spoiling. They contain acids, which keep the pH low enough to stop the growth of bacteria.

Two ingredients in pancakes are buttermilk (an acid) and baking soda (a base). The two mix, producing a salt, water and carbon dioxide. The carbon dioxide bubbles get bigger when the mixture is heated, causing the mixture to rise.

Cooking with acids and bases

Body basics

Your body uses alkalis to neutralise acidity in various parts of the digestive and excretory system. When food leaves the stomach and passes into the duodenum, it has some of the hydrochloric acid from the stomach mixed into it. This could burn the tender interiors of the intestines further down if not neutralised. To counter this acidity, a greenish-yellow alkali called **bile** is secreted by the gall bladder, entering the duodenum through the bile ducts. Here the bile mixes with the mush of partly digested food and acid and makes it neutral.

In the swim

When chlorine is added to a swimming pool, it reacts with the water to produce hypochlorous acid. This acid kills bacteria and algae, keeping the pool water safe for swimming. All the chemicals in a swimming pool, when combined, need to have a pH in the range of 7.2–7.8 for a clean, hygienic pool and safe swimming.

If the pH falls below 7.2, the micro-organisms will still be killed but the swimmers will get red and stinging eyes, and the water may become corrosive and damage pool fittings. A base such as sodium carbonate (soda ash) or sodium bicarbonate (bicarbonate of soda) would have to be added to neutralise the excess acid.

If the pH rises above 7.8, bacteria and algae will grow and the water will be unfit for swimming. To reduce the pH, an acid such as sodium hydrogen sulfate would have to be added to neutralise the excess base.

HOW ABOUT THAT!

The tastebuds on the human tongue can detect five different kinds of taste — sweet, sour, bitter, salty and savoury. Powdered sherbet stimulates three of these as it is made of sour tasting citric acid, bitter bicarbonate of soda and sweet powdered sugar. When you eat sherbet, the water from your saliva allows a chemical reaction to happen between the citric acid and the bicarbonate of soda to produce carbon dioxide gas that gives a fizzing sensation that you can feel in your mouth.

ACTIVITIES

REMEMBER

1 **Distinguish** between acids and bases.

2 What common reaction do some acids and bases have when they come into contact with solid substances?

3 **Describe** the difference between a base and an alkali.

4 **Identify** which type of substance has a pH value:
(a) less than 7 (b) more than 7 (c) equal to 7.

5 **Recall** why the chemical reaction between an acid and a base is called neutralisation.

6 **Recall** what is produced in all neutralisation reactions.

7 **Identify** which acid can be found in your stomach.

8 **Explain** how an antacid relieves the pain of indigestion.

9 **Explain** how self-raising flour helps cakes rise.

USING DATA

10 A pH meter is used to measure the pH of 5 different substances. The results are as shown in the table below:

Substance	A	B	C	D	E
pH valve	6.0	12	3.0	7.0	8.0

(a) **Identify** which substance is most likely to be: (i) orange juice (ii) milk.

(b) **Identify** which substance could be: (i) a weak base (ii) pure water (iii) vinegar (iv) a strong base.

(c) **Identify** which two of the substances you would expect to be the most corrosive.

THINK

11 A burning feeling in your stomach is often due to the juices in your stomach becoming too acidic. The treatment for this 'indigestion' problem is to take an antacid tablet. Antacid tablets contain a base, which neutralises the excess acid and relieves the pain. When you take an antacid tablet, would you expect the pH value in your stomach to increase or decrease? **Explain** your answer.

12 A stinging-nettle plant may contain an acid that is injected into your skin when you touch it. **Describe** how you could show that the plant does contain an acid.

13 Write a balanced chemical equation to describe the chemical reaction between hydrochloric acid and calcium carbonate.

14 Why is it that the acids in the food and drink you consume do not damage your stomach?

15 When you add buttermilk (an acid) to baking soda (a base) in a mixing bowl, does the pH increase or decrease? **Explain** your answer.

INVESTIGATE

16 On the packet of one brand of baking soda, there is a claim that you can deodorise your entire house by sprinkling it on your carpets and leaving it for a few minutes before vacuuming.
(a) Suggest how the baking soda could have this deodorising effect.
(b) **Investigate** other claims made on baking soda packaging and design some experiments to test one or more of the claims.

eBook*plus* work sheet 7.3 Acids, bases and neutralisation

Salty substances

Tea, coffee, shampoo and baking powder are all examples of household chemicals designed to be used when dissolved in water. Because water is a good **solvent**, it is able to dissolve many other substances. Chemicals which are dissolved are called **solutes.** A **solution** is a mixture of a solute and the solvent in which it is dissolved. Many reactions involving acids and bases occur in water. These reactions are said to occur 'in solution'.

In the previous section, we have seen how the excess stomach acid that causes indigestion can be treated by neutralising it with a base. The reaction between the acid and a base produces water and a salt in a process called neutralisation. Note that we say *a* salt. You may think that there is only one kind of salt — the kind that you put on your fish and chips. In reality, the word salt describes any product of the neutralisation process that contains a positive metal ion and a negative non-metal ion bonded together. So while common table salt (sodium chloride) is indeed a salt in the chemical sense so too are potassium chloride, sodium sulfate and magnesium nitrate.

The base sodium hydrogen carbonate, commonly known as 'bi-carb', is a component of baking powder. It has the formula $NaHCO_3$ and contains the hydrogen carbonate ion HCO_3^-. When bases containing this ion react with acids, the gas carbon dioxide is produced as well as a salt and water. As we saw in the previous section, when hydrochloric acid and bi-carb are mixed together, the following reaction takes place:

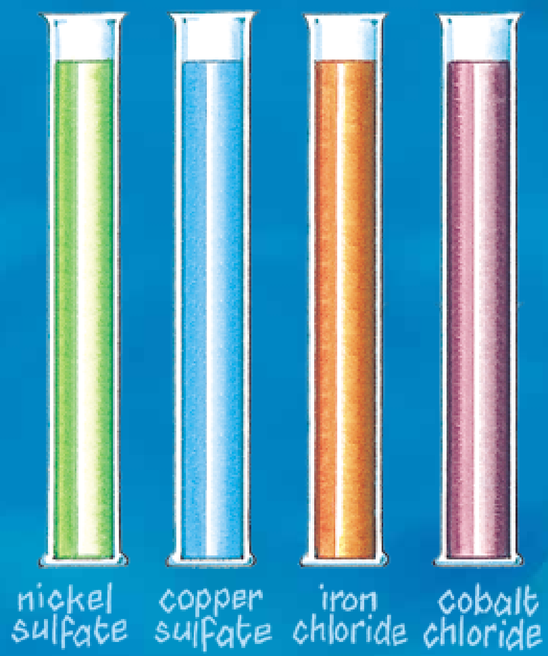
$$NaHCO_3(s) + HCl(aq) \rightarrow NaCl(aq) + CO_2(g) + H_2O(l)$$

base + acid → salt + carbon dioxide + water

The products are sodium chloride, carbon dioxide and water. In the above reaction, the salt formed was a metal chloride, because it contains the chloride ion (Cl^-) from the hydrochloric acid. Neutralisation reactions between many different acids and bases are possible; therefore, it is possible to produce many different salts. The names of all these salts are related to the bases and acids from which they are formed. Some of these are summarised in the table below.

HOW ABOUT THAT!

Many salts are brightly coloured and many are highly poisonous and not at all suitable for sprinkling on your fish and chips! Salts containing copper ions are usually blue, those containing nickel are pale green, those containing iron can be green or orange and cobalt salts are pink.



Not all salts can be sprinkled on your fish and chips!

Base	Acid	Negative ion present in salt	Salt
Sodium hydroxide	Sulfuric acid	Sulfate SO_4^{2-}	Sodium sulfate
Magnesium oxide	Hydrochloric acid	Chloride Cl^-	Magnesium chloride
Sodium oxide	Acetic acid	Acetate CH_3COO^-	Sodium acetate
Copper(II) oxide	Nitric acid	Nitrate NO_3^-	Copper(II) nitrate

INVESTIGATION 7.5

Pass the salt!

AIM To produce salt from a neutralisation reaction

You will need:

safety glasses and laboratory coat
50 mL burette
retort stand, bosshead and clamp
tripod and gauze mat
Bunsen burner, heatproof mat and matches
20 mL pipette
100 mL conical flask
pipette bulb
white tile
dropping bottle of phenolphthalein indicator
wire shaped into a loop with a handle
small funnel
1 mol L^{-1} hydrochloric acid solution
1 mol L^{-1} sodium hydroxide solution
evaporating dish
silver nitrate solution in a dropping bottle
sample of sodium chloride
test tube

CAUTION Wear safety glasses and a laboratory coat.

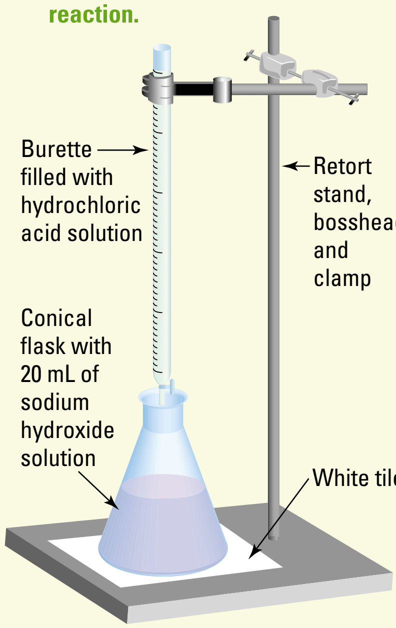
- Rinse the burette with the hydrochloric acid solution and then, using the funnel, fill the burette with the hydrochloric acid solution.
- Rinse the pipette with sodium hydroxide solution using the pipette bulb.

CAUTION Never pipette using your mouth.

- Set up the equipment as shown in the diagram on the right. Use the pipette and bulb to transfer 20 mL of the sodium hydroxide solution into the conical flask.
- Add a few drops of phenolphthalein indicator to the sodium hydroxide.
- Add the acid from the burette carefully until the pink colour of the indicator disappears. The colour change indicates that the neutralisation reaction is complete.
- Pour the contents of the flask into an evaporating dish. Heat the dish with the Bunsen burner and gently evaporate the water. Be careful — splattering may occur.
- When the water has nearly evaporated, turn off the Bunsen burner and allow the dish to cool and the remaining water to evaporate without further heating.
- Test the white crystals for the presence of sodium ions by placing a few crystals on a wire loop and heating in a Bunsen burner flame. Compare this flame colour with a known sample of sodium chloride. Record your observations.
- Test for the presence of chloride ions by dissolving a few crystals in half a test tube of water and adding a few drops of silver nitrate. A white cloudiness indicates that chloride ions are present. Record your observations.

DISCUSSION

1. **Comment on the information that the flame and silver nitrate tests provided. What conclusion can you draw?**
2. **Write a word equation for the neutralisation reaction.**
3. **Write a balanced equation, using formulae, for the neutralisation reaction.**
4. **Design a test to show that water was the other product of the reaction.**



ACTIVITIES

REMEMBER

1. **Explain** the meaning of the words 'solute', 'solvent' and 'solution'.
2. **Recall** what the expression 'reaction in solution' means.
3. **Define** the term 'salt'.
4. **Recall** the products of a reaction between an acid and a base that contains a hydrogen carbonate ion.

THINK

Use the table on page 241 and the table below to answer the following questions.

5. Write balanced chemical equations for the following reactions.
   (a) Solid sodium bicarbonate and sulfuric acid react to form a sodium sulfate solution, carbon dioxide and water.
   (b) Solid potassium hydroxide and hydrochloric acid react to form a solution of potassium chloride and water.
   (c) Solid copper oxide reacts with sulfuric acid to form a solution of copper sulfate and water.
6. **Identify** the salts that would form from the reaction between:
   (a) magnesium hydroxide and hydrochloric acid
   (b) potassium hydroxide and acetic acid
   (c) sodium carbonate and sulfuric acid.

Common laboratory bases	Base formula
Sodium hydroxide	$NaOH$
Copper hydroxide	$Cu(OH)_2$
Magnesium oxide	MgO
Potassium hydroxide	KOH
Magnesium hydroxide	$Mg(OH)_2$
Sodium carbonate	Na_2CO_3
Sodium bicarbonate	$NaHCO_3$

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work sheet 7.4 Neutralisation and salts

Precipitation reactions

When table salt (sodium chloride) is dissolved in water to form an aqueous solution, it seems to 'disappear'. The ions in the salt no longer bond together as a large lattice of positive and negative ions like they do as a solid. The sodium ions and the chloride ions separate when they dissolve in water. Because the separated ions are so small, we cannot see them. However, when we evaporate water from the solution we find that sodium chloride crystals are left behind. The dissolving of sodium chloride in water can be represented by the equation:

$$NaCl(s) \xrightarrow{(H_2O)} Na^+(aq) + Cl^-(aq)$$

Ions in aqueous solutions are therefore separate entities and are able to react independently.

Presto precipitation!

When colourless lead nitrate solution and colourless potassium iodide solution are added together, a brilliant yellow solid is formed. Where does this solid come from?

When two solutions containing dissolved ions are mixed together, these ions are able to come into contact with each other. Oppositely charged ions attract. In some cases, the attraction is strong enough to form ionic bonds and hence a new ionic compound. Some of these compounds are insoluble (unable to dissolve in water) and so a solid forms. This solid is called a **precipitate.** Chemical reactions in which precipitates form are called **precipitation reactions**.

Ionic compounds dissolve in water to varying degrees. Some are said to be soluble, others slightly soluble and others insoluble. The box here outlines some handy rules for predicting if a compound is soluble or not.

Soluble or not?

1. All compounds containing either the Na^+, NH_4^+, K^+ or NO_3^- ion will dissolve in water. Compounds containing these ions never form precipitates.
   Example: This rule tells us that NaCl, NH_4Cl, K_2SO_4 and $AgNO_3$ are all soluble in water and therefore do not form precipitates.
2. Compounds containing the Cl^-, Br^- and I^- ions are soluble, except when they contain the Ag^+, Pb^{2+} or Hg^{2+} ions.
   Example: This rule tells us that $FeCl_3$, $ZnBr_2$ and AlI_3 are soluble, but that AgCl, $HgBr_2$ and PbI_2 are not soluble.
3. Compounds containing the SO_4^{2-} ion are soluble, except for $BaSO_4$, $PbSO_4$ and $CaSO_4$.
   Example: This rule tells us that $ZnSO_4$ will dissolve, but $BaSO_4$ will form a precipitate.
4. Compounds containing CO_3^{2-} and PO_4^{3-} are insoluble except when they contain the ions Na^+, NH_4^+ or K^+.
   Example: This rule tells us that $BaCO_3$ will form an insoluble precipitate, but Na_2CO_3 will not.
5. Compounds containing OH– are insoluble, unless they contain the ions Na+, NH4+ or K+.
   Example: This rule tells us that $Zn(OH)_2$ will form a precipitate, but NaOH will not.
6. Some compounds are slightly soluble. These include $Ca(OH)_2$, $PbCl_2$, $PbBr_2$, $CaSO_4$ and Ag_2SO_4.

Changing partners

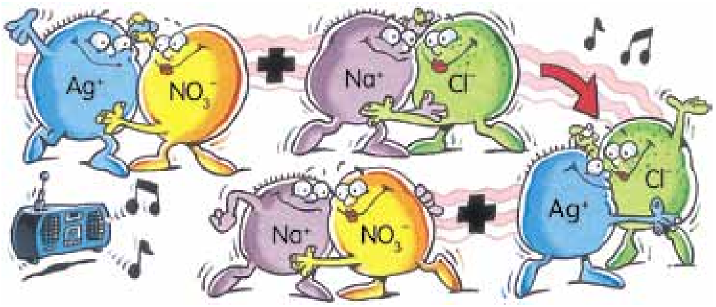
Another example of a precipitation reaction is the one between silver nitrate solution and sodium chloride solution. When these two clear, colourless solutions are added together, the contents of the test tube become cloudy, indicating that a precipitate has formed. If the tube is allowed to stand for a while, the solid settles to the bottom and we can see that a clear solution is also present. The products of the reaction are insoluble solid silver chloride (the precipitate) and sodium nitrate (not visible because it is soluble in water). This reaction can be represented by the equations:

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

$$AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$$

The box above indicates that silver nitrate, sodium chloride and sodium nitrate all dissolve in water. Therefore, they have the symbol '(aq)'. Silver chloride does not dissolve in water so it has the symbol '(s)' to indicate that it is solid. The precipitation of silver chloride can be used as evidence of the presence of silver ions in an unknown solution.

The equation on page 243 shows that the ions in the reactants have changed 'partners'. The silver ion is paired with the chloride ion on the product side of the reaction and the sodium ion is paired with the nitrate ion. The opposite is the case on the reactant side of the equation. A positive ion can pair up only with a negative ion because oppositely charged ions are attracted to each other. When writing the formula of any new compound, the positive ion is always written first.



Ions sometimes change 'partners' when a chemical reaction takes place.

INVESTIGATION 7.6

Will it precipitate?

AIM To investigate precipitation reactions

You will need:

5 semi-micro test tubes and test-tube rack
a white tile
a black tile
safety glasses
dropping bottles of the following solutions: copper sulfate, sodium chloride, silver nitrate, cobalt chloride, sodium hydroxide and potassium iodide

CAUTION *Wear safety glasses.*

- Place 10 drops of copper sulfate solution in each test tube.
- Add 10 drops of sodium chloride to the first test tube, 10 drops of silver nitrate to the second, and so on until each tube contains copper sulfate solution and one other solution. Hold a black or white tile behind the test tube if necessary to detect the presence of a precipitate.
- If there is a reaction, record your observations in a table.
- Tip the residues into a waste bottle. Wash out the test tubes thoroughly and this time place 10 drops of sodium chloride in each of the test tubes. Again add one of the other solutions to each of the test tubes (but not copper sulfate as this combination has already been tested). Record your observations in your table.
- Repeat until all possible pairs of solutions have been tested.

DISCUSSION

1. **Write word equations for each of the pairs that reacted to form a precipitate.**
2. **Use formulae to write equations for each of the pairs that reacted to form a precipitate.**
3. **You could have predicted which pairs of solutions would form a precipitate using the box of solubility rules on the previous page. Check to see if the rules match your results.**

ACTIVITIES

REMEMBER

1 **Define** the term 'precipitate'.

2 Why don't all reactions between salts produce precipitates?

3 **Explain** how a reaction between colourless substances can produce a coloured precipitate.

THINK

4 Are all precipitates insoluble in water? **Explain**.

5 Write an equation for the reaction that occurs when the salt, copper sulfate, dissolves in water.

6 **Identify** which two of the following compounds will be soluble in water.
(a) $NaNO_3$
(b) KI
(c) PbI_2
(d) $Zn(OH)_2$

7 **Identify** which of the following compounds will be insoluble in water.
(a) $CuCO_3$
(b) AgI
(c) NaCl
(d) $Mg(OH)_2$

8 Write down the possible combinations of ions when the following solutions are mixed together:
(a) sodium chloride and copper sulfate
(b) sodium hydroxide and copper sulfate
(c) lead nitrate and sodium hydroxide
(d) potassium iodide and sodium carbonate.

9 For each of the reactions listed in question 8, **identify** the precipitate that would form. If you believe that no precipitate would form, write 'no precipitate'.

INVESTIGATE

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10 To find out more about precipitation and other reactions use the **Introduction to reactions** weblink in your eBookPLUS.

work sheet — 7.5 Precipitation

Metal displacement reactions

The word *displace* means 'to push out of place', and that pretty much sums up displacement reactions. Displacement reactions occur where one metal pushes another out of a compound and takes its place.

For example, when a strip of magnesium metal is added to lead sulfate, the magnesium pushes the lead out of the sulfate compound and joins with it. As a result, the products of the reaction are magnesium sulfate and lead (which precipitates out). The equation for this reaction is:

$$Mg(s) + PbSO_4(aq) \rightarrow MgSO_4(aq) + Pb(s)$$

Some metals are more likely to take part in displacement reactions than others. Gold and silver for example, very rarely take part in displacement reactions.

A measure of how likely a particular metal is to take part in a displacement reaction is its **reactivity**.

Some metal elements, like potassium, lithium and sodium, are so reactive that they have to be stored under oil to stop them from reacting with substances in the air. The more reactive metals are never found in a pure form naturally — they are always found in the form of a metal compound. Unreactive metals such as gold, on the other hand, are most likely found as a pure metal nugget or seam.

The activity series

Lithium, sodium and potassium are stored under oil so that they can't react with the air. These substances react quickly and explosively when in contact with water in the air.

The activity series places the metallic elements in decreasing order of reactivity. This series is shown below.

In order for a metal to displace another in a compound, it must be higher on the activity series. So, if we put magnesium metal in a copper sulfate solution, the magnesium will displace the copper to form magnesium sulfate. However, if we were to put a strip of copper into a magnesium sulfate solution, no displacement would occur — they'd pretty much just sit there and look at each other!

In order for a metal to react with acid and release hydrogen gas, the metal must be before hydrogen in the activity series.

The reactivity of metals can be investigated by observing the reactions of metals with acids. When a metal reacts with hydrochloric acid, it reacts according to the equation:

metal + hydrochloric acid → salt + hydrogen gas

In these reactions, electrons are transferred away from the metal atoms to the hydrogen in the acid, forming positive metal ions and hydrogen gas. These are referred to as **redox reactions** (we will look at what a redox reaction is in more depth in the next section). As the metal has displaced the hydrogen from the acid, these reactions are also displacement reactions.

HOW ABOUT THAT!

Gold is so unreactive that it can be placed in concentrated hydrochloric, nitric or sulfuric acid without effect. No single acid is able to dissolve gold. However, there is one substance that will. Alchemists called it *aqua regia* (*'royal water'*) and it is made up from a mixture of concentrated hydrochloric and nitric acid. It can also dissolve platinum, yet the metals osmium, tantalum and iridium remain unaffected by it!

Li	K	Na	Ca	Mg	Al	Mn	Cr	Zn	Fe	Ni	Sn	Pb	H	Cu	Hg	Ag	Au	Pt

Decreasing reactivity →

INVESTIGATION 7.7

Measuring the reactivity of metals

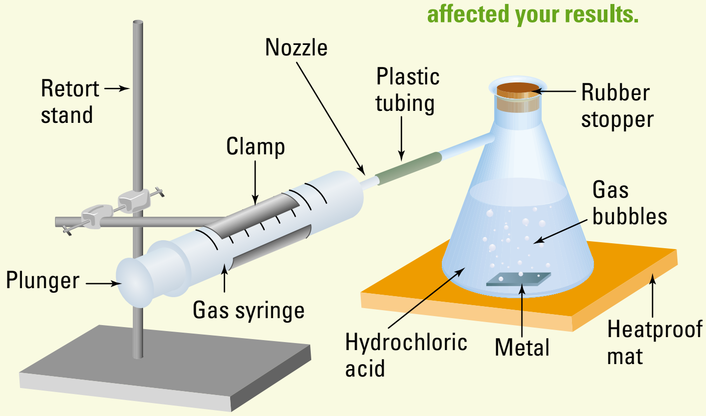
AIM To construct a metal activity series

You will need:

safety glasses and laboratory coat
steel wool
1 cm × 4 cm pieces (or equivalent amount) of copper, zinc, aluminium, iron and magnesium
gas syringe
retort stand, bosshead and clamp
heatproof mat
distilled water
1 cm diameter × 4 cm long piece of plastic tubing
1 mol L^{-1} hydrochloric acid
100 mL measuring cylinder
250 mL side-arm conical flask
rubber stopper to fit conical flask
stopwatch or clock with a second hand

CAUTION *Wear safety glasses.*

- Before starting this investigation, read all of the instructions and construct a table suitable for recording your measurements.
- Use the steel wool to polish each of the samples of metal.
- Mount the gas syringe in the clamp as shown in the diagram at right. Your teacher will tell you if the syringe needs to be lubricated. Push the plunger in fully and attach the plastic tubing to the nozzle.
- Carefully measure out 50 mL of hydrochloric acid and pour it into the conical flask.
- Connect the free end of the plastic tubing onto the arm of the flask.
- Prepare to start timing. Carefully add one of the pieces of metal to the flask and quickly seal it with the rubber stopper. Start timing as soon as the metal is placed in the flask.
- Record (in your table) the volume of gas in the syringe every 30 seconds until gas is no longer produced, the syringe is full or 10 minutes has passed, whichever occurs first.
- Repeat this procedure with the other metal samples. Rinse the conical flask with distilled water each time before repeating the procedure.
- When you have completed your measurements, plot a graph of the results on a single set of axes. Your graph should show how the volume of gas for each sample changes with time. That is, plot volume of gas on the vertical axis and time on the horizontal axis.



DISCUSSION

1. **Use your graph to list the five metal elements in order of reactivity.**
2. **Write a word equation for the reaction of each of the metal elements with the acid. If no reaction occurred, write 'no reaction'.**
3. **Some of the variables in this investigation were not well controlled. List them and explain how the lack of control may have affected your results.**

ACTIVITIES

REMEMBER

1 **Recall** why pure gold is more likely to be found in the ground than pure sodium.

2 **Describe** the activity series of metals.

3 **Identify** the gas that is always formed when a metal reacts with an acid.

THINK

4 Balance these chemical equations that describe the reaction between acids and metals

(a) $Zn(s) + HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$

(b) $Na(s) + HCl(aq) \rightarrow NaCl(aq) + H_2(g)$

5 You have a choice of making a piece of jewellery from either zinc or nickel. Which metal would you choose? Justify your choice.

6 The elements gold, silver, copper and iron were all discovered more than 6000 years ago. Yet, the elements potassium, sodium and calcium were not discovered until 1808. Suggest a reason for this.

INVESTIGATE

7 Find out at least four different ways of treating iron (other than mixing it with other elements to make an alloy) to prevent or reduce its corrosion.

8 Design and carry out an experiment that investigates the reactivity of alloys, such as stainless steel and brass. Compare these results with those obtained for the metal elements. Present the findings of your investigation as a poster to display in your classroom.

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work sheet — 7.6 Metal displacement

Reactions everywhere!

In a world where countless chemical reactions take place, it is helpful to classify the reactions. They can be classified according to whether they release or absorb energy, and can also be grouped together according to the nature of the reactants, the nature of the products, the way in which charged particles in atoms rearrange themselves or even the number of reactants. Because there are different ways of classifying chemical reactions, any one reaction can fall into several different groups.

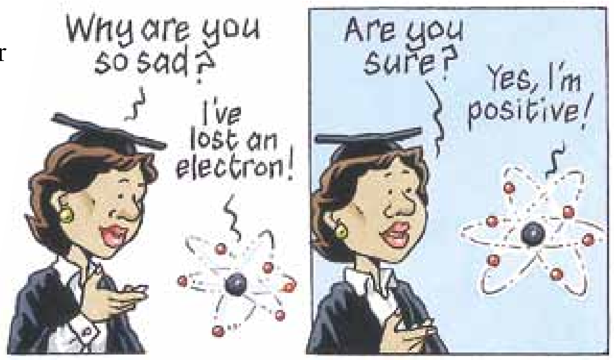
Explaining reactions — redox

In many chemical reactions, electrons are either completely or partially moved from one atom, ion or molecule to another. This process is known as **electron transfer**. Chemical reactions that involve electron transfer are called redox reactions. Redox reactions are extremely important in industry and in our everyday lives.

A redox reaction is really two reactions occurring simultaneously. In the electron transfer process, one reactant loses electrons and another gains electrons. Loss of electrons is known as **oxidation**. Gain of electrons is called **reduction**. Oxidation and reduction always occur together, thus the two words are combined to form the word redox, which is used to describe reactions where electrons are transferred.

The mnemonic OIL RIG may help you to remember these processes: Oxidation Is Loss, Reduction Is Gain.

The corrosion, displacement, combustion and combination reactions described on the next few pages are examples of redox reactions. Oxidation and reduction can be clearly seen in the reaction that occurs when zinc corrodes (see following page).



Corrosion reactions

Corrosion is a chemical reaction that occurs between a metal and substances occurring in the air or water around it that causes the metal to be eaten away over time. The tarnish that appears on silver cutlery, the milky green verdigris that builds up on copper and bronze statues and the dull grey layer that coats bare aluminium window frames are all examples of corrosion, as is the rusting of iron.

Corrosion of zinc

If you look at a sheet of galvanised iron, you will notice that it does not have a shiny metallic surface. Galvanised iron has been coated with a layer of

The rusting of this shed is an example of corrosion.

zinc metal. The zinc prevents the iron underneath from reacting with the oxygen and water in the air and rusting. Instead, it is the zinc that corrodes, reacting with oxygen to form a dull layer of zinc oxide on the surface of the metal. The chemical equation for this reaction is:

$$2Zn(s) + O_2(g) \rightarrow 2ZnO(s)$$

Now, let's consider this from the redox point of view. When the zinc corrodes, electrons are transferred from the zinc atoms to the oxygen molecules, causing the formation of positive zinc ions and negative oxide ions. These oppositely charged ions attract and bond together to form the ionic compound zinc oxide:

2 zinc atoms	+	1 oxygen molecule	→	$2Zn^{2+}$ ions	+	$2O^{2-}$ ions
$2Zn^{2+}$	+	O_2	→	$2ZnO$		

In this reaction, zinc atoms lose electrons, thus zinc is oxidised. Oxygen molecules gain electrons, thus oxygen is reduced. Remember that oxidation and reduction always occur together.

Combustion reactions

Combustion reactions are those in which a substance reacts with oxygen and heat is released. Examples of combustion reactions include the burning of petrol in a motorcycle engine, wax vapour in a candle flame and natural gas in a kitchen stove. In each of these cases **hydrocarbons** (compounds containing only the elements carbon and hydrogen) combine with oxygen in the air to form carbon dioxide gas and water vapour.

Combustion of methane

The chemical equation for the burning of methane is:

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

methane + oxygen → carbon dioxide + water

In this redox reaction, electron transfer is not complete. The reactants are molecules and the products are also molecules. In each molecule, electrons are shared by the atoms. However, the oxygen atoms in the products attract the electrons more strongly than the carbon and hydrogen atoms. Therefore, the shared electrons spend more time close to the oxygen atoms. The electrons have been partially transferred to the oxygen atom. Thus, oxygen is reduced and the carbon in methane is oxidised.

Combination reactions

A **combination reaction** (also known as a **synthesis reaction**) is one in which two reactants produce a single product usually accompanied by a release of energy in the form of heat and/or light. One spectacular example of this type of reaction is that between magnesium and oxygen. As the magnesium metal burns in air it produces a brilliant white light. The equation for this combination reaction is:

$$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$$

magnesium + oxygen → magnesium oxide

INVESTIGATION 7.8

Decomposing powder

AIM To carry out a quantitative investigation of a decomposition reaction

You will need:

laboratory coat and safety glasses
zinc carbonate powder
spatula
Bunsen burner, heatproof mat and matches
large pyrex test tube and test-tube rack
test-tube holder
an electronic balance
marking pen
stereo microscope
Petri dish

CAUTION *Wear safety glasses and laboratory coat.*

- Place two spatulas of zinc carbonate powder in the test tube. Weigh the test tube and record the mass.
- Mark the level of the powder in the test tube with the marking pen.
- Heat the test tube gently in a blue Bunsen burner flame for 5 to 10 minutes.

CAUTION *Make sure the test tube is not pointing at anyone.*

- While heating the test tube, hold a lit match at the mouth of the tube. Record your observations.
- Allow the test tube to cool down. Note any change in the level of powder and then reweigh the test tube. Record the mass.
- Place small amounts of zinc carbonate and the powder from the test tube in the Petri dish. Examine them using a stereo microscope. Record your observations.

DISCUSSION

1. Which gas was given off during the reaction?
2. Explain any change that occurred in the mass.
3. Write word and formula equations for the reaction.

In this reaction, electrons are transferred from the atoms in the magnesium metal to the oxygen atoms in the oxygen molecule. This forms positive metal ions and negative oxide ions. These ions are attracted to each other due to their opposite charges and form the white, ionic, solid magnesium oxide. Magnesium, which loses electrons, is oxidised, and oxygen is reduced.

$$2Mg(s) + O_2(g) \rightarrow 2Mg^{2+} + 2O^{2-} \rightarrow 2MgO(s)$$

Decomposition reactions

In decomposition reactions one single compound breaks down into two or more simpler chemicals. This is pretty much the opposite of a combination reaction. An example of this is the decomposition of zinc carbonate which is represented by the equation:

$$ZnCO_3(s) \rightarrow ZnO(s) + CO_2(g)$$

zinc carbonate → zinc oxide + carbon dioxide

Displacement reactions

A displacement reaction occurs when a more reactive metal displaces a less reactive metal and takes its place in a compound. As we saw in the previous section, silver has a low reactivity and is easily 'pushed out' of compounds by more reactive metals. When copper reacts with silver nitrate, the more reactive copper displaces the silver atom to form copper nitrate, leaving the silver to precipitate out as a solid:

$$Cu(s) + 2AgNO_3(aq) \rightarrow 2Ag(s) + Cu(NO_3)_2(aq)$$

copper + silver nitrate → silver + copper nitrate

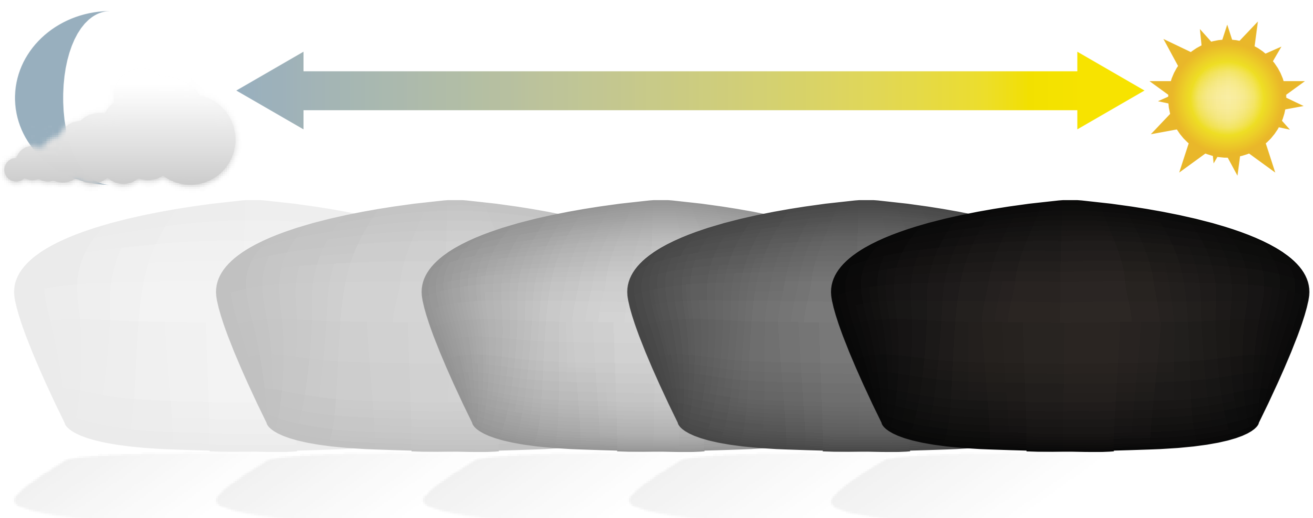
In this reaction, electrons are transferred from the copper atoms to the silver ions. Silver ions (Ag^+) in the solution gain electrons to form atoms of solid silver. Thus, silver ions are reduced. Copper atoms (Cu(s)) lose electrons, forming copper ions (Cu^{2+} (aq)), which dissolve into a solution. The formation of copper ions changes the colour of the solution from colourless to blue. The copper atoms are oxidised. The nitrate ion is not involved in the electron transfer. Ions that are not involved in ion transfer are referred to as **spectator ions**.

Redox in the light and the shade

People who wear glasses often don't want to bother with swapping over to sunglasses when they go outside. **Photochromic** glasses solve the problem by darkening as the wearer moves from indoors into bright sunshine. They lighten again when the wearer moves back into an area of low light. Plastic photochromic glasses use organic material which darkens the lenses when exposed to ultraviolet light. Glass photochromic glasses work due to the presence of silver chloride (AgCl) crystals in the glass. When a wearer is in the sunshine, ultraviolet light is absorbed by the silver chloride crystals and a redox reaction occurs. Electrons are transferred from the chloride ion to the silver ion according to the equation:

$$Ag^+ \text{ (ion)} + Cl^- \text{ (ion)} \rightarrow Ag \text{ (atom)} + Cl \text{ (atom)}$$

Silver particles then form in the glass, darkening the lens so that visible light is absorbed and reflected.



Photochromic lenses darken in the presence of bright light then become clear when the ambient light becomes dim because of a series of redox reactions.

The fading of the dark glass is more complicated. The chlorine atoms are very reactive. To stop them reacting with the silver atoms and reversing the process too quickly, singly charged copper ions are dissolved in the molten glass during the manufacturing process. These ions react with the chlorine atoms to form chloride ions and doubly charged copper ions in the reaction:

$$Cu^+ + Cl \rightarrow Cu^{2+} + Cl^-$$

When the glasses are no longer in the sunlight, the doubly charged copper ions accept an electron from the silver atom. The silver ion re-forms and the dark lens becomes light again:

$$Cu^{2+} + Ag \rightarrow Cu^+ + Ag^+$$

Reactions with a zap!

The chemical reactions that produce electrical energy in electric cells (more commonly known as batteries) are redox reactions. In electric cells, electrons are transferred from one reactant to another through the wires that make up the electric circuit. This is very useful because the moving electrons can provide the energy to operate our appliances. Thus, chemical energy from the redox reaction is converted to electrical energy. The reactants in the cells are not in direct contact with each other. In an ordinary carbon battery or dry cell, the reactants are separated by a paste that allows the movement of electric charge. The electrons flow from one reactant at the negative electrode, through the electric circuit to the other reactant at the positive electrode. Chemical products are formed at both electrodes.

ACTIVITIES

REMEMBER

1 **Construct** a table similar to the one below and use it to summarise each of the groups of reactions discussed on pages 247–9. List one example of a reaction for each group.

Reaction type	Description	Example

2 **Recall** what all redox reactions have in common.

3 **Define** the term 'oxidation'.

4 **Define** the term 'reduction'.

5 **Explain** where the word redox came from.

6 Consider the reaction:

$$2Zn(s) + O_2(g) \rightarrow 2ZnO(s)$$

(a) **Identify** which reactant the electrons are being transferred from.
(b) **Identify** which reactant the electrons are transferred to.

7 Give an example of a redox reaction where the electron transfer is not complete.

THINK

Refer to the tables on pages 188–9 of chapter 5 and the table on page 241 of this chapter to answer question 8.

8 Write a balanced equation using formulae for the following reactions:

(a) **copper metal** + **zinc sulfate solution** → **zinc metal** + **copper sulphate solution**
(b) **sodium metal** + **oxygen gas** → **solid sodium oxide**
(c) **carbon monoxide gas** + **oxygen gas** → **carbon dioxide gas**
(d) hydrogen peroxide (H_2O_2) solution decomposes to form hydrogen gas and oxygen gas.

9 **Identify** the type of each of the reactions in question 8 (remember, a reaction may be more than one type!).

10 **Explain** how it can be said that the reaction between magnesium and oxygen is four reactions in one: 'a combustion reaction', 'a combination reaction', 'a redox reaction' and 'an exothermic reaction'.

INVESTIGATE

11 Find out what a Daniell cell is and how redox can be applied to the way in which it produces electric current.

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12 Test your ability to identify different types of reactions by completing the **Time Out 'Reactions'** interactivity in your eBookPLUS. **int-0759**

13 Use the **Chemical reactions** weblinks in your eBookPLUS to learn more about the different types of chemical reactions.

work sheet 7.7 Corrosion and combustion

Seeing the light!

It's a scene that you always see in crime shows on the TV. The forensic scientists spray a mysterious chemical over a seemingly clean carpet, they turn out the lights and suddenly, you see a glowing spatter pattern and drag marks appearing. A footprint is outlined where it stepped through the blood that has since been cleaned away. This is the scene of a crime! So what is this mysterious glowing chemical? Why doesn't it glow while it's being sprayed on the carpet?

Glowing in the dark

While combustion reactions such as the burning of wood may produce light as a result of the heat produced, some reactions release light with very little heat being produced.

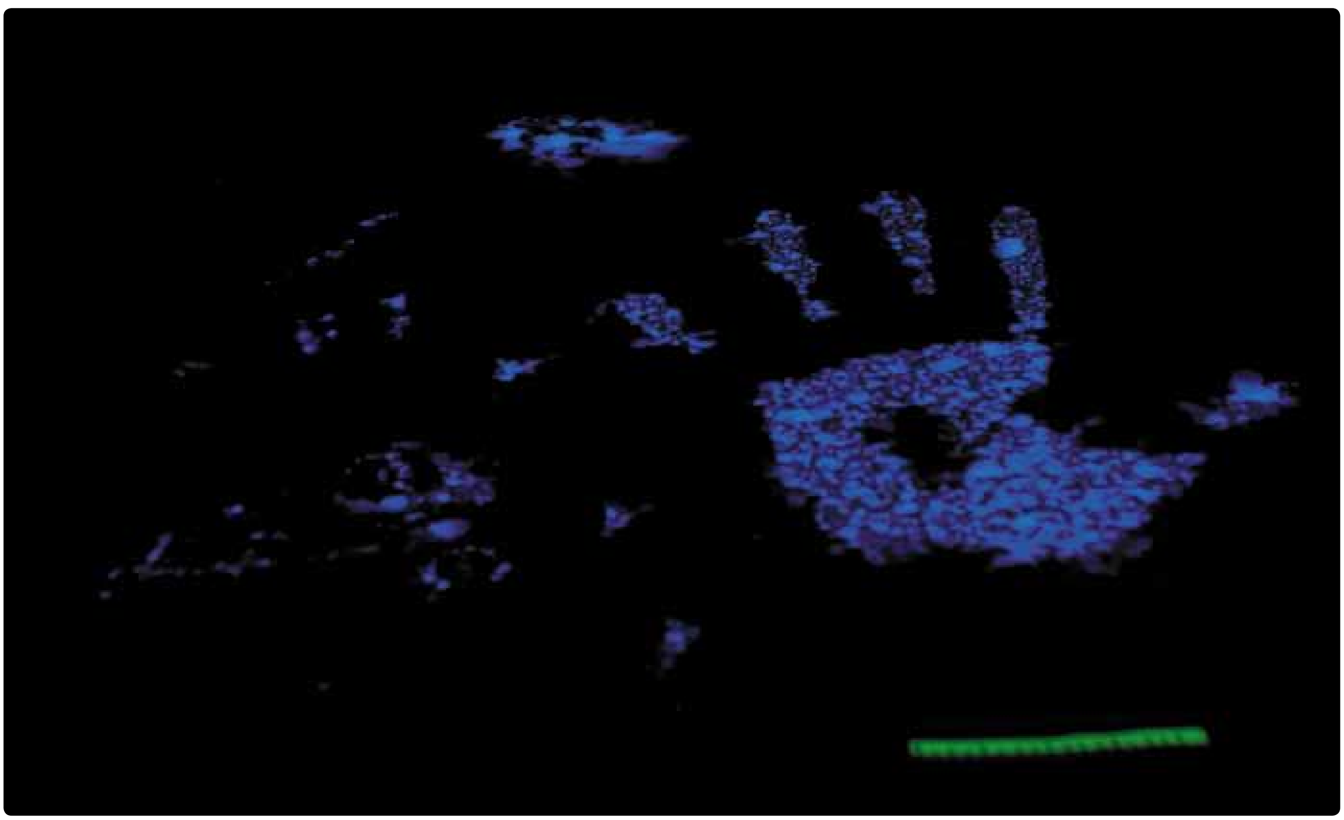
Chemiluminescence is light that is produced as a result of a chemical reaction. This light is called 'cold light' because very little heat is produced during these reactions. Luminol and glow sticks both produce chemiluminescence as the result of specific reactions.

Crime scene chemicals

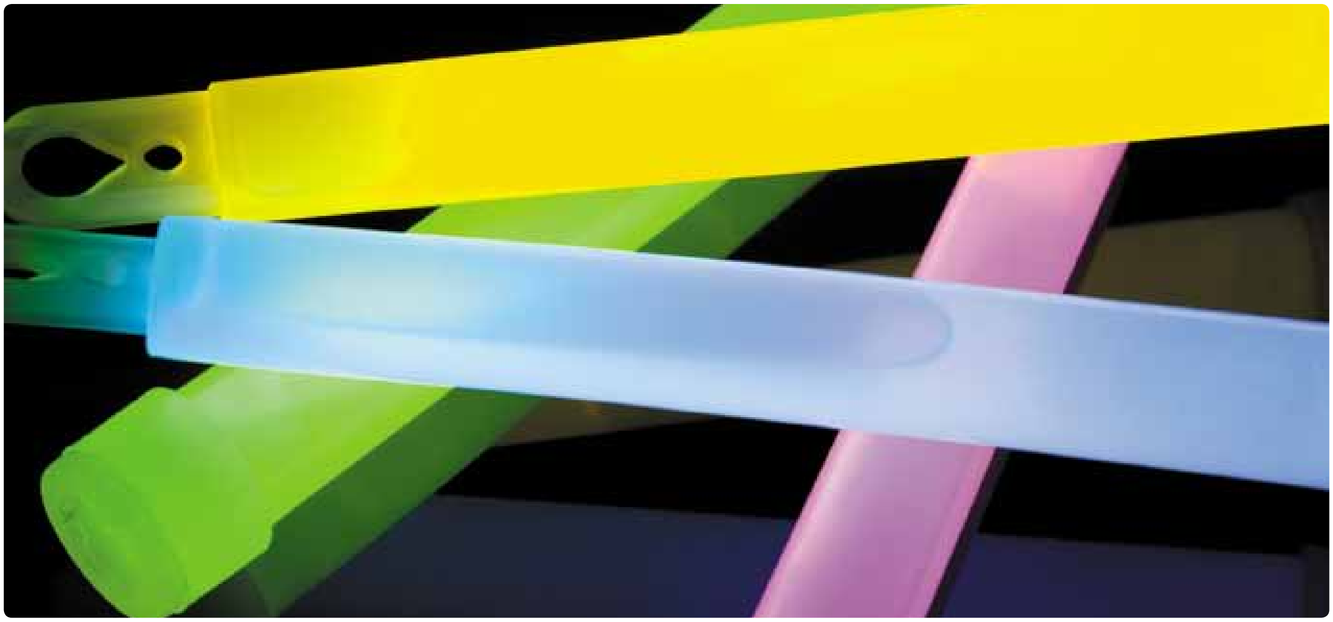
Luminol is a chemical that reacts with hydrogen peroxide to produce aminophthalate and light. This reaction is normally slow and the light produced is quite weak. However, some substances such as the iron found in blood act as **catalysts** for the reaction.

$$\text{Luminol} + \text{hydrogen peroxide} \xrightarrow{\text{Iron}} \text{aminophthalate} + \text{light}$$

Crime scene technicians use a mixture of Luminol and hydrogen peroxide to detect traces of blood even if none is visible to the naked eye. The area is sprayed evenly with the mixture and, where even tiny amounts of blood may be present, the iron in the blood causes the reaction between the Luminol and the hydrogen peroxide to speed up. The reaction produces a blue glow which can be seen if the room is made dark. This glow lasts about 30 to 40 seconds allowing the blood patterns to be photographed and recorded for later comparison with the lit crime scene.



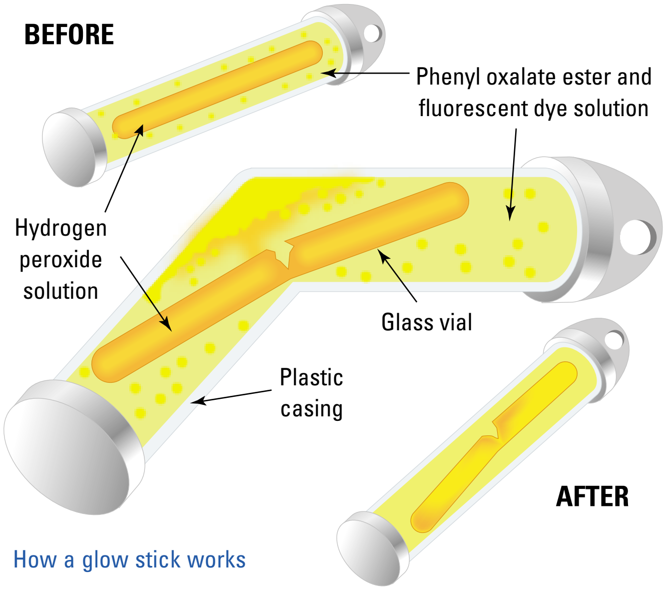
Revealed in light — the scene of the crime



Glow sticks

Glow sticks are a big party favourite and they also rely on chemiluminescence for their light. A glow stick is made up of two tubes, one inside the other. The small inner tube is made of glass and contains a solution of hydrogen peroxide while the larger flexible outer tube holds a chemical called phenyl oxalate ester and a coloured fluorescent dye.

When you bend the glow stick, the inner tube is cracked open, allowing the hydrogen peroxide to react with the ester. This reaction produces a chemical called phenol, and carbon dioxide and energy which stimulates the dye to produce light without producing heat. The colour of the light produced depends upon the colour of the fluorescent dye present. Eventually, when all of the reactants are used up, the glow stick stops glowing.



How a glow stick works

That natural glow

Bioluminescence is the light produced by living organisms such as fireflies, glow worms and many deep ocean creatures. This light is caused by chemical reactions within the body of the organism itself, so bioluminescence can be thought of as a form of chemiluminescence.

The key reaction in bioluminescence occurs between a natural pigment called **luciferin** and oxygen. Another chemical, **luciferase**, acts as a catalyst in the reaction, allowing light to be produced from the organism's body. Different organisms produce different colours of bioluminescence according to the type of luciferin that they have.

Mimicking bioluminescence

The production of cool light by fireflies has been used as a model for the development of chemiluminescent materials. Although the production of light by chemiluminescence has been possible for some time, commercial applications were often not developed because the reactions were relatively inefficient. The firefly is able to produce light very efficiently by the chemical reactions in the cells of its abdomen. However, in recent years chemical research has uncovered new chemiluminescent reactions and more efficient reactions have been developed. This has enabled the commercial production of chemiluminescent items and the use of chemiluminescence techniques in scientific research.

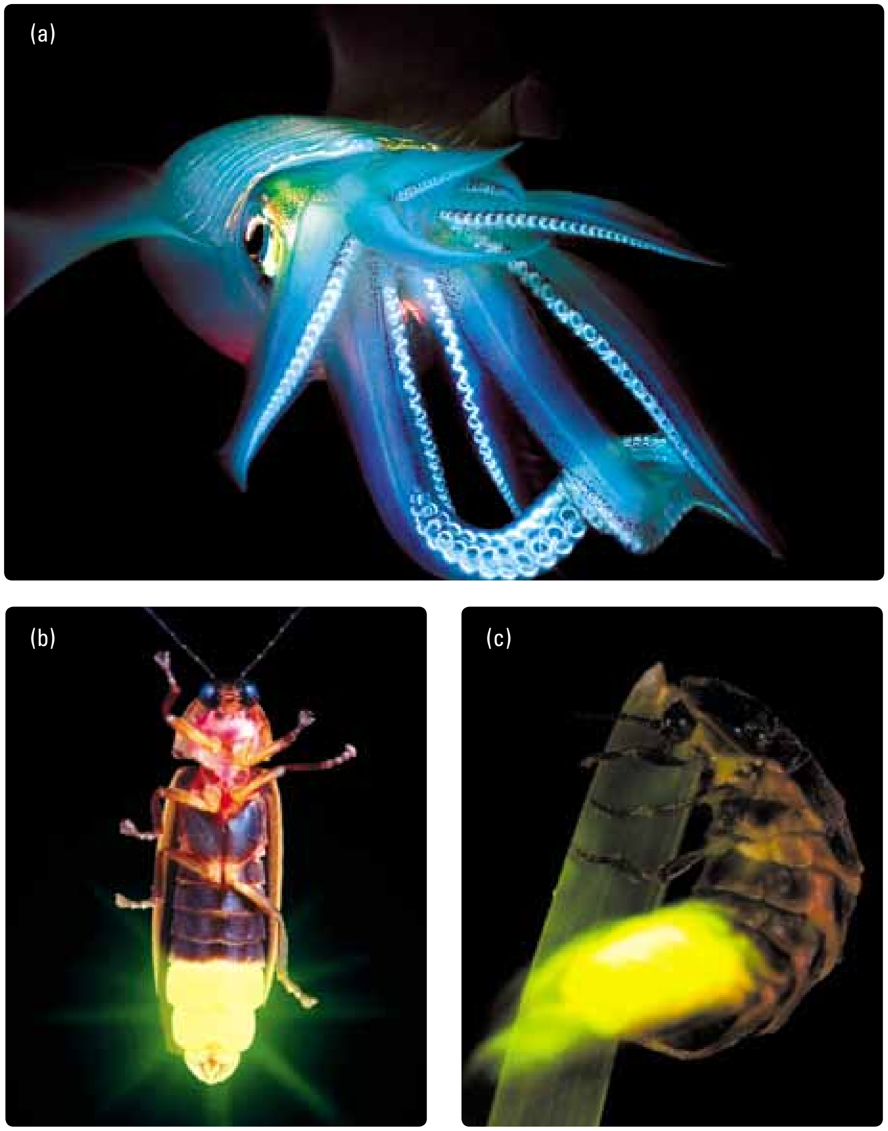
Using chemiluminescence and bioluminescence

The reactions that occur in chemiluminescence and bioluminescence have been adapted for use in scientific research, medicine, ecology, hygiene and food quality control.

Bioluminescence is used when testing for tuberculosis to determine the most suitable antibiotic to be given to the patient. Scientists have used gene transfer technology to insert the firefly's gene for making luciferase enzymes into bacteria from the tuberculosis patient. These bioluminescent bacteria are then tested for their resistance to different antibiotics.

The effectiveness of the antibiotics can be easily determined by the amount of bioluminescence remaining.

Bioluminescent bacteria have also been used to test for mercury pollution in water. No doubt in the future many more uses will be found for chemiluminescence and bioluminescence.



Many different creatures exhibit bioluminescence including (a) squid (b) fireflies and (c) glow worms.

ACTIVITIES

REMEMBER

1 What is chemiluminescence?

2 Give two examples of chemiluminescent reactions.

3 What are the (a) reactants (b) products in the reaction in a glow stick?

4 How does chemiluminescence differ from bioluminescence?

5 **Describe** at least two ways in which chemiluminescence differs from the luminescence provided by a light bulb.

6 Draw a diagram to explain how light is produced in a chemiluminescent light stick.

THINK

7 Most of the organisms found in the deep ocean display bioluminescence. Why do you think this is the case?

8 When magnesium burns in oxygen, it produces an intense white light. Is this an example of chemiluminescence? **Explain** your answer.

9 Glow sticks aren't just good for producing party light. Give two other occasions in which glow sticks are put to use.

INVESTIGATE

10 One of the chemicals involved in bioluminescence is called luciferase. Find out how it got this name.

11 In crime shows, you sometimes see blood and other body fluid traces being detected by technicians shining an ultraviolet light (or 'black' light) on a surface. **Investigate** why this works and what its limitations are.

12 Use the internet and other sources to further **investigate** luminol and answer the following questions.
   (a) **Identify** the chemicals that luminol is made from.
   (b) **Identify** other substances, apart from blood, that will make luminol glow.

Taking care with chemicals

Many of the chemicals used in industry, medicine, schools, universities and homes can be hazardous to your health. The hazards come about because these chemicals can react with parts of your body — inside or out. Apart from the dangers to your own health, chemicals can, as a result of their properties or their reactions with common substances such as water and air, cause great damage to property and the environment.

Laws exist, at both national and state level, to ensure that people using harmful chemicals are informed about how to handle and use them safely. For this purpose, harmful chemicals are placed within one or both of the groups known as **dangerous goods** or **hazardous substances**.

Dangerous goods

Chemicals in the dangerous goods group are those that could be dangerous to people, property or the environment. Most dangerous goods are grouped into one of nine classes, according to the greatest immediate risk they present. Some of the classes are divided into subclasses. Dangerous goods must be identified with the appropriate dangerous goods sign on their labels. The table on the next page lists nine of the classes and subclasses, along with their respective label signs. These signs must, by law, be prominently placed and clearly visible. Their design also includes a symbol for the type of hazard they present.

Outside these nine classes, there are two other groups of dangerous goods.

1. Goods too dangerous to be transported (GTDTBT).
2. Combustible liquids (C1), which includes liquids that are not as easily ignited as flammable liquids, but which will ignite at temperatures below their boiling point.

Hazardous substances

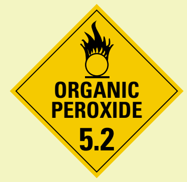
Chemicals in the hazardous substances group are those that have an effect on human health. The effect may be immediate, such as poisoning and burning, or long term, such as liver disease or cancer. Hazardous substances can enter the body in a number of ways. They can be inhaled, absorbed through the skin, ingested (swallowed) or injected.

Hazardous substances are identified on their labels by a 'signal word' providing a warning about the substance, or the word 'hazardous' printed in red. 'Signal words' include 'dangerous poison', 'poison', 'warning' and 'caution'. Labels of hazardous substances also include:

- information about the risks of the substance
- directions for use
- safety information
- first aid instructions and emergency procedures.

If the substance is also in the dangerous goods group, the label will also include the appropriate diamond sign showing its class.

Classes and subclasses of dangerous goods

Class	Description	Sign
Class 1	Explosive substances or articles used to produce explosions	
Class 2.1	Flammable gases: gases that ignite in air if in contact with a source of ignition such as a spark or flame	
Class 2.2	Non-flammable, non-toxic gases: these gases may cause suffocation	
Class 2.3	Toxic gases: gases likely to cause death, serious illness or injury if inhaled	
Class 3	Flammable liquids: liquids with vapours that can ignite on contact with air at temperatures below 60°C	
Class 4.1	Flammable solids: solids that are easily ignited by a source of ignition such as a spark or flame	
Class 4.2	Substances liable to spontaneous combustion: solids that can ignite without an external source of ignition	
Class 4.3	Substances that emit flammable or toxic gases on contact with water	
Class 5.1	Oxidising agents: substances that may contribute to the combustion of other substances, increasing the risk of fire	
Class 5.2	Organic peroxides: substances that undergo exothermic decomposition reactions	
Class 6.1	Toxic substances: chemicals likely to cause death, serious illness or injury if swallowed, inhaled or brought into contact with skin	
Class 6.2	Infectious substances: substances containing micro-organisms likely to cause diseases in humans or animals	
Class 7	Radioactive material	
Class 8	Corrosive substances: substances that corrode metals or cause injury by reacting on contact with living tissue	
Class 9	Miscellaneous dangerous goods and articles: dangerous substances and objects that do not belong to the other classes	

Keeping you informed

All employers are required by law to make sure that their employees are fully informed about the chemicals in the workplace that are classified as dangerous goods and/or hazardous substances.

A list of such chemicals stored or used in the workplace must be kept, along with a copy of the chemical's **MSDS** (**material safety data sheet**). Chemical suppliers are required to provide a MSDS for each of the hazardous substances or dangerous goods that they supply. In turn, employers are required to make the MSDS accessible to employees who are exposed to the chemicals.

A MSDS is likely to consist of several A4 pages and many can be downloaded directly from the internet. The information on a MSDS should include:

- the ingredients of the product
- the date of issue — an up-to-date MSDS should be no more than 5 years old
- information about health hazards and first aid instructions
- precautions that need to be taken when using the product
- information about storage and safe handling of the product.

Assessing risk

A risk assessment identifies the potential hazards of an experiment and gives protective measures to minimise the risk. Before any experiment involving chemicals is conducted in your school laboratory, a risk assessment is carried out. The form of a risk assessment varies from school to school, but will always contain:

- a summary of the experiment
- a list of the risks and safety precautions for each chemical
- information about whether the chemical is classified as a hazardous substance or dangerous good
- a list of protective measures to be taken. These might include the use of a fume hood and/or the wearing of safety glasses or other protective items.
- first aid information.

Most of the information used in a risk assessment is obtained from the MSDS for each of the chemicals used. The date on the MSDS used for each chemical must be stated to ensure that the risk assessment is up to date.

Risk assessment sheets in schools are usually completed and signed by a qualified science teacher or laboratory technician. Your science teacher is required to carefully read the risk assessment sheet before allowing an experiment involving chemicals to commence.

ACTIVITIES

REMEMBER

1 **Describe** what the chemicals listed as dangerous goods have in common.

2 If a chemical in the dangerous goods group is explosive, toxic and corrosive, how is it decided to which class the chemical belongs?

3 **Recall** what the chemicals listed as hazardous substances have in common.

4 **Identify** four 'signal words' used on the labels of hazardous substances.

5 What is a MSDS and what should it include?

6 Where can employers obtain a MSDS for hazardous substances and dangerous goods?

7 Whose responsibility is it to make sure that people have access to an MSDS for each of the hazardous chemicals and dangerous goods that they store or use?

THINK

8 **Describe** what characteristics chemicals listed as both dangerous goods and hazardous substances have in common.

9 **Distinguish** between flammable liquids (Dangerous goods, Class 3) and explosive liquids (Dangerous goods, Class 1).

10 **Explain** the difference between the purposes of a MSDS and a risk assessment sheet.

11 **Explain** why every chemical used in a laboratory (including water) should be considered to be a health hazard.

INVESTIGATE

12 Many chemical suppliers provide access to MSDSs online. Use the internet to search for a MSDS on hydrochloric acid and use it to answer the following questions.

(a) **Identify** some alternative names for hydrochloric acid.

(b) What are the health hazards of hydrochloric acid?

(c) **Describe** the first aid treatment recommended if hydrochloric acid: (i) is ingested (swallowed) (ii) is inhaled (iii) makes contact with an eye (iv) makes contact with the skin.

(d) **Summarise** the recommendations made for the storage of hydrochloric acid.

LOOKING BACK

FOCUS activity

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Devise a presentation in the media of your choice to explain the chemical reactions that occur in one day of your daily life.

Access more details about focus activities for this chapter in your eBookPLUS.

doc-10654

1. French chemist Antoine-Laurent Lavoisier provided evidence that led to the development of the Law of Conservation of Mass and the Law of Constant Proportions.
   (a) Use the Law of Conservation of Mass to **explain** why it is incorrect to say that when a candle burns it disappears.
   (b) **Recall** the Law of Constant Proportions.

2. In an experiment to test the effect of the amount of liver on the breakdown of hydrogen peroxide, the following results were obtained.

Mass of liver (g)	Volume of oxygen released (cm^3)
0.5	2.5
1.0	5.1
2.0	9.8
2.5	11.5

   (a) Write a word equation for the reaction occurring in this experiment.
   (b) Use formulae to write an equation for this chemical reaction.
   (c) Graph these results on graph paper.
   (d) **Describe** what the graph shows about the effect of the liver on the rate of this reaction.

3. When an aqueous solution of barium hydroxide reacts with an aqueous solution of ammonium hydroxide, the temperature of the products becomes low enough to freeze water.
   (a) **Define** the term 'aqueous solution'.
   (b) Is this an example of an exothermic or endothermic chemical reaction? **Explain** your answer.
   (c) **Deduce** where the energy transferred to or from the reactants goes.

4. Write balanced equations using formulae for the following reactions.
   (a) **aluminium metal + oxygen gas → solid aluminium oxide**
   (b) **potassium metal + oxygen gas → solid potassium oxide**
   (c) **solid carbon + oxygen gas → carbon dioxide gas**
   (d) **solid copper carbonate → solid copper oxide + carbon dioxide gas**
   (e) **iron metal + sulfur powder (S_8) → solid iron sulfide (FeS_2)**
   (f) **copper sulphate solution + zinc metal → copper metal + copper sulphate solution**
   (g) **copper(II) sulfate solution + sodium hydroxide solution → solid copper(II) hydroxide + sodium sulfate solution**
   (h) **solid magnesium hydroxide + hydrochloric acid → magnesium chloride + water**

5. The two reactants in the chemical reaction taking place in the test tube shown below are aqueous solutions. There is enough evidence in the photograph to identify the type of chemical reaction taking place.

   (a) **Identify** what type of chemical reaction this is.
   (b) **Explain** what evidence in the photograph identifies the type of reaction.

6. **Identify** the reaction type (displacement, decomposition, precipitation, combustion or neutralisation) for each of the reactions in question 5.

7. Which of the reactions in question 5 are redox reactions?

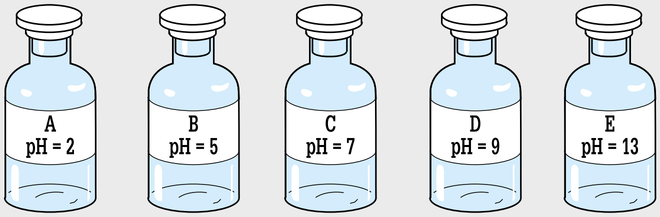
8. Many chemicals are classified as dangerous goods and/or hazardous substances.
   (a) **Describe** the differences between these two categories of chemicals.
   (b) **Recall** what these two categories of chemicals have in common.

9. What is a MSDS and what is it used for?

10. **Explain** why you are more likely to find pure gold than pure copper in the ground.

11. **Predict** the salts that would result from the neutralisation reaction between:
   (a) magnesium oxide and hydrochloric acid
   (b) copper(II) oxide and sulfuric acid
   (c) sodium hydroxide and acetic acid
   (d) sodium oxide and nitric acid.

12 The liquids in the bottles shown below are labelled with their pH. **Identify** which of the bottles is most likely to contain:
(a) distilled water
(b) a strong acid
(c) a strong base
(d) vinegar
(e) bathroom surface cleaner.



13 An experiment was carried out to measure how long it took for equivalent amounts of different metals to dissolve in 100 mL of hydrochloric acid. The results obtained are shown below.

Metal	Time taken for the metal to dissolve (min)
Iron	1.0
Magnesium	3.0
Tin	2.5
Aluminium	3.5
Nickel	2.0
Zinc	1.5

Unfortunately, the person recording the data on the computer accidentally changed the order of the metals recorded in the table. Using the known activity series table on page 245, redraw the table so that the correct metal is matched with the correct time.

14 Early scientists mixed dilute hydrochloric acid with some calcium carbonate in a beaker and carefully weighed the beaker and its contents. After the bubbling had stopped, they discovered that there was a drop in mass. They used this experiment to state that the Law of Conservation of Mass was not true. Do you agree or disagree with the statement? **Explain**.

15 Currently, we use chemical equations to describe chemical reactions. How else could you describe a chemical reaction? Use an example to **explain** your new method.

TEST YOURSELF

1 What is the only reliable evidence indicating that a chemical reaction has taken place?
A A change in temperature
B A change in state
C Formation of a new substance
D Disappearance of one or more reactants (1 mark)

2 Which of the following are products of the reaction between silver nitrate and sodium chloride?
A Silver nitrate and sodium chloride
B Nitrogen chloride and silver sodium
C Do not react so there will be no products
D Silver chloride and sodium nitrate (1 mark)

3 The most reactive metal in the activity series is
A lithium.
B potassium.
C platinum.
D iron. (1 mark)

4 Which of the following is a balanced equation?

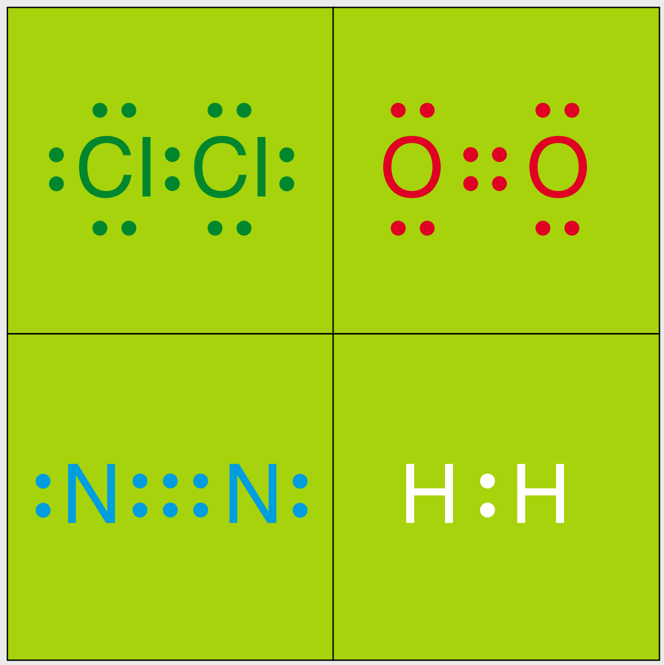
A $Na + 2Cl \rightarrow 2NaCl$

B $MgO + 2HCl \rightarrow MgCl{+}H_2$

C $MgO + 2HCl \rightarrow MgCl_2 + H_2O$

D $2Na + Cl \rightarrow 2NaCl$ (1 mark)

5 The electron dot diagrams for chlorine gas, hydrogen gas and oxygen gas are shown below. The number of electrons in the outer shells of the elements that form these gases are 7, 1 and 6 respectively.



(a) Draw electron dot diagrams for
(i) hydrochloric acid (HCl)
(ii) water (H_2O)
(iii) carbon tetrachloride (CCl_4)
(iv) carbon dioxide (CO_2).
(b) Look at the electron dot diagram for nitrogen.
(i) **Identify** how many electrons it has in its outer shell.
(ii) Draw an electron dot diagram for ammonia (NH_3).
(4 marks)

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work sheets
7.8 Chemical reactions puzzle
7.9 Chemical reactions summary

STUDY CHECKLIST

COMMUNICATING CHEMISTRY

- ☐ write word equations to **describe** common reactions between chemicals 7.1
- ☐ **distinguish** between reactants and products 7.1, 7.2
- ☐ **define** the terms 'exothermic' and 'endothermic' with regard to reactions 7.3
- ☐ write balanced chemical equations using chemical formulae to **describe** important chemical reactions between substances 7.2
- ☐ **recall** the Law of Conservation of Mass and the Law of Constant Proportions 7.1

USING CHEMICALS

- ☐ **distinguish** between an acid and a base 7.4
- ☐ **predict** the products that are formed as the result of neutralisation reactions 7.4
- ☐ **describe** the role of chemical indicators 7.4
- ☐ **recall** common uses of both acids and bases in everyday life 7.4
- ☐ **define** the term 'chemical salt' 7.4, 7.5
- ☐ **distinguish** between hazardous and dangerous chemicals 7.10
- ☐ **recall** the purpose of a risk assessment and a MSDS 7.10

CHEMICAL REACTIONS

- ☐ **describe** the reactants and the products in a variety of neutralisation, combustion, corrosion, precipitation, decomposition and metal displacement reactions 7.4, 7.5, 7.6, 7.7, 7.8
- ☐ **distinguish** between combustion and corrosion 7.8
- ☐ **describe** what redox is and explain how the term can be applied to the chemical reactions already studied 7.8
- ☐ **explain** how the formation of a precipitate as a result of a chemical reaction may be predicted 7.6
- ☐ **define** the term 'chemiluminescence' 7.9

APPLICATIONS AND USES OF SCIENCE

- ☐ **define** the terms 'chemiluminescence' and 'bioluminescence' and recall some of their uses 7.10

SUMMARY

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Digital documents

Individual pathways

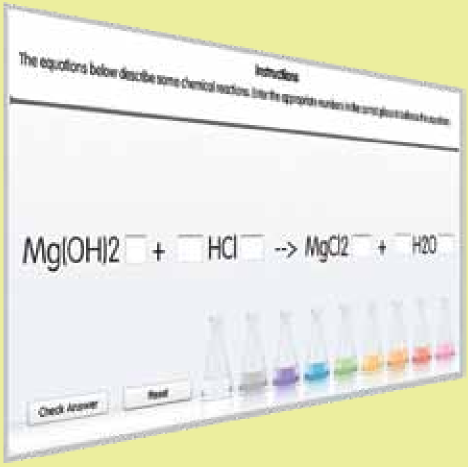
Activity 7.1	Activity 7.2	Activity 7.3
Revising chemical reactions	Investigating chemical reactions	Investigating chemical reactions further
doc-10655	doc-10656	doc-10657

Interactivities

Time Out 'Reactions'
Use this exciting interactivity to test your ability to classify different types of reactions before time runs out.

Searchlight ID: int-0759

Checking for balance
Use this interactivity to test your ability to balance chemical equations.



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projects*plus*

ChemQuiz!

SEARCHLIGHT ID: PRO-0107

Scenario

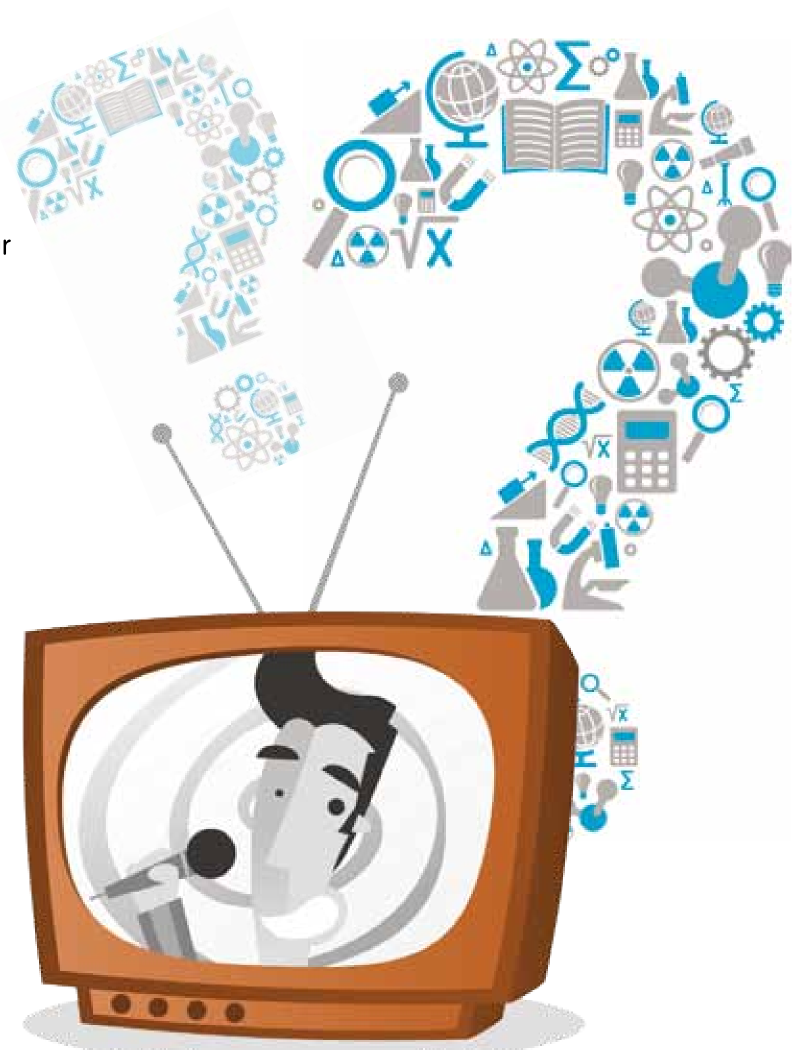
You only have to have a glance at any page of your TV guide to see that Australians young and old love a good quiz show. Whether it's *Hot Seat, Jeopardy, Spit it Out* or *It's Academic*, programs with a quiz show format rate consistently well. While the idea of watching someone answer questions seems like an odd form of entertainment, psychologists theorise that their popularity arises from a combination of a desire to learn new information and a form of competition — after all, who hasn't watched a quiz show and yelled the answers at the screen? In recent educational studies, the use of quiz game formats as a teaching tool in the classroom is gaining support.

The Brain Mine is a company that specialises in educational resources for use in Science classrooms. On the basis of these educational studies of quiz games, they have decided that they would like to add a computer-based chemistry quiz show that teachers could purchase and run in their classrooms as a fun and effective way of improving student knowledge. As product developers at The Brain Mine, it is up to you and your team to make this happen! You and your team are going to develop *ChemQuiz*, a chemistry-based quiz show in which the class teacher will act as the show host, groups of students will be the contestants and the questions (which pop up on a computer screen so that the contestants can see them) are based on chemistry skills.

Your task

Using PowerPoint, you will create a series of question screens for a quiz show that should run for about ten minutes. For each question screen, the show host must be able to reveal the correct response after a contestant has given their answer. The question screens should be entertaining and eye-catching, and should also be easily readable by the contestants and the show host (who will read the questions out as they appear).

You will need to give a demonstration of your *ChemQuiz* show with one of your group acting as the show host (the role that would normally be taken by the teacher). The show host will need to explain the rules of the quiz show at the start. The contestants will be your fellow students (preferably not those in your group, who will already know the answers!).

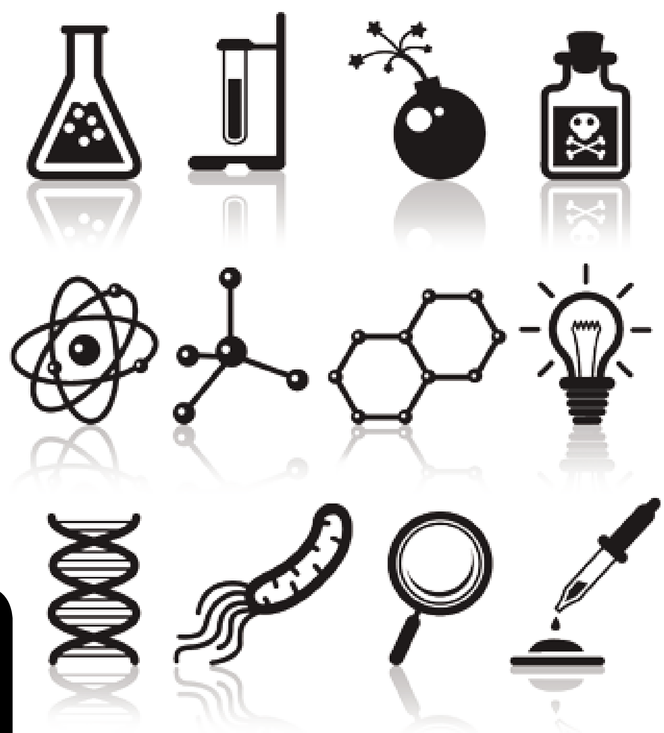


Process

- Open the ProjectsPLUS application for this chapter located in your eBookPLUS. Watch the introductory video lesson and then click the 'Start Project' button to set up your project group. You can complete this project individually or invite other members of your class to form a group. Save your settings and the project will be launched.
- Navigate to your Research Forum. Here you will find a number of tabs labelled with research topic headings that will help you organise your question ideas. You may add new research topics if you want.
- Start your research. Make notes of ideas that you can use when creating your quiz questions, such as interesting chemical facts, balancing equations, remembering chemical symbols and names, determining the products of a chemical reaction and so on — remember that the audience and contestants for *ChemQuiz* will be Year 9 or Year 10 students. Enter your ideas, questions and answers (which MUST be correct) as articles under your topic headings in the Research Forum. You can view and comment on other group members' articles and rate or correct the information that they have entered. When your research is complete, print out your Research Report to hand in to your teacher.
- Visit your Media Centre and download the PowerPoint gameshow template to help you build your question screens for the quiz. Your Media Centre also includes images and video clips that you may find useful to add to your question screens where appropriate. There are also sound effects that you may like to add to indicate that the contestant has given correct or incorrect answers.
- Use your questions, answers and PowerPoint to create your *ChemQuiz* gameshow!

SUGGESTED SOFTWARE

- ProjectsPLUS
- PowerPoint
- Internet access



MEDIA CENTRE

Your Media Centre contains:

- a sample rule book
- a selection of useful weblinks
- a selection of images
- an assessment rubric.

Your ProjectsPLUS application is available in this chapter's Student Resources tab inside your eBookPLUS. Visit www.jacplus.com.au to locate your digital resources.