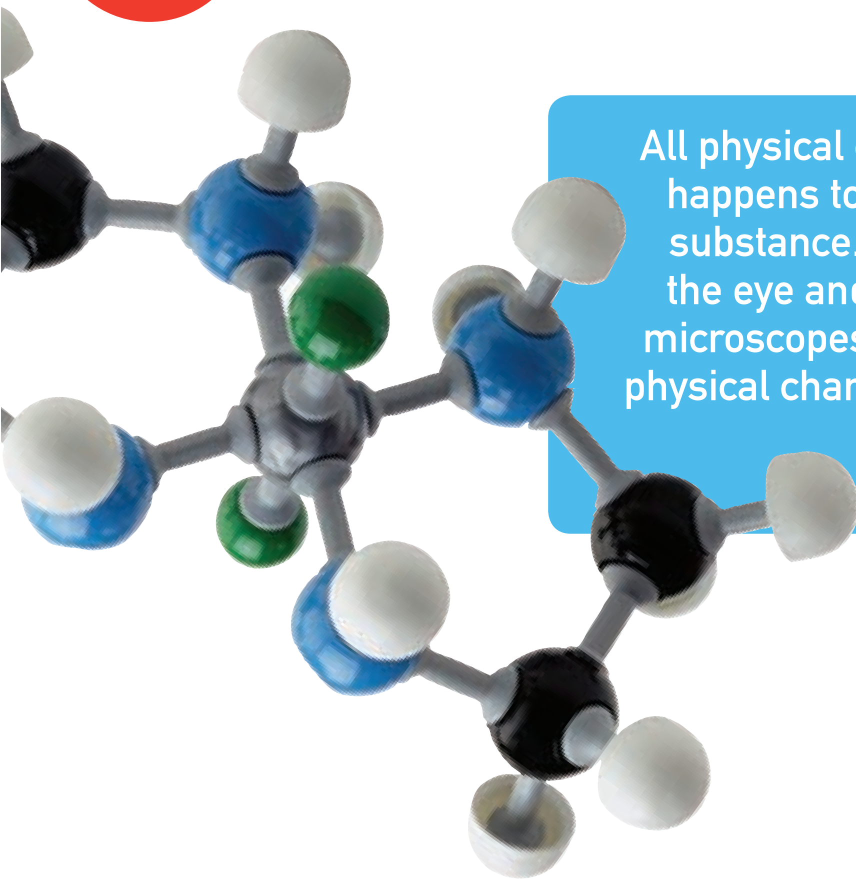
6.2 Understanding physical change



All physical changes occur because of what happens to the particles making up a substance. These particles are invisible to the eye and even the most powerful optical microscopes. Therefore, to understand physical change, scientists use models to help explain and understand their observations.

Understanding physical change

Most physical changes can be understood by using the **particle model**, as shown in Figure 6.2.1. The particle model is a simplified representation of solids, liquids and gases. It can explain some but not all of their properties.

The particle model assumes that all forms of matter (solids, liquids and gases) are made up of invisible, ball-like particles that are:

- hard, incompressible (not able to be squashed) and indivisible (unable to be split)
- attracted to each other
- constantly moving.

In solids, the attraction between particles binds them tightly and rigidly together. For this reason, solids are incompressible and hold their own shape. Although the particles in solids are fixed in position, they vibrate on the spot. These vibrations increase with temperature.

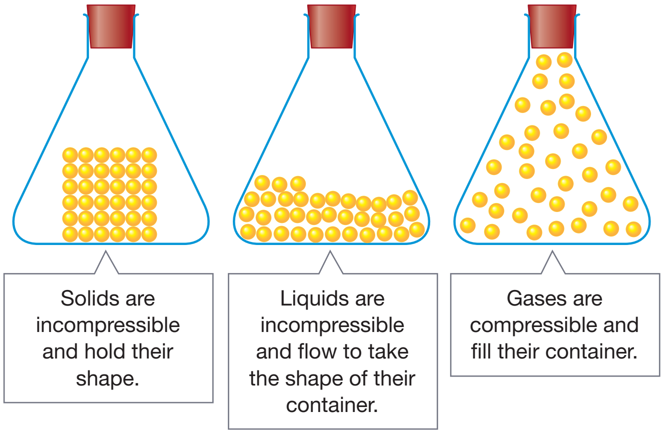


Figure 6.2.1 The physical properties of solids, liquids and gases can be explained by the particle model.

In liquids, the particles are packed tightly together, which makes liquids incompressible. However, the particles in a liquid are not stuck rigidly to each other so liquids flow to take on the shape of their container. Particles in a liquid vibrate but can also move freely throughout the liquid. The particles move faster as the temperature increases.

In gases, the particles are not stuck to each other at all and there are large spaces between gas particles. For this reason, gases are highly compressible. The particles are free to move anywhere within their container and move in straight lines until they collide with another gas particle or the side of the container. This is why gases will always fill their container.

Expansion, contraction and pressure

The physical changes of expansion and contraction can be understood by looking at how the motion of particles changes with temperature. In solids, the particles vibrate in fixed positions. As the temperature increases, so too do the vibrations—pushing the particles further apart and causing the solid to expand (get larger) as shown in Figure 6.2.2. When the solid is cooled, the reverse happens. The particles vibrate less, allowing them to be packed more tightly and the solid contracts (shrinks).

The particles in liquids also vibrate more when they are heated. This causes liquids to expand when heated and contract when cooled. However, in liquids the particles are free to move, so liquids tend to expand and contract more than solids.

Gases will always expand or contract to fill their container. However, as the temperature of a gas is increased, the particles travel faster. This means they hit the sides of their container more frequently and with more force. If the container is flexible (such as the balloon in Figure 6.2.3) then increasing the temperature of the gas will cause the balloon to expand. If the container is rigid, such as a glass bottle, then heating the gas will cause the pressure inside the container to increase.

The reverse is also true. If a gas is cooled, then the particles travel slower and hit the sides of their container less frequently and with less force. This results in lower pressure. If the gas is cooled in a flexible container, then the container will contract. A rigid container will not contract but the pressure inside the container will be reduced.

1 p219

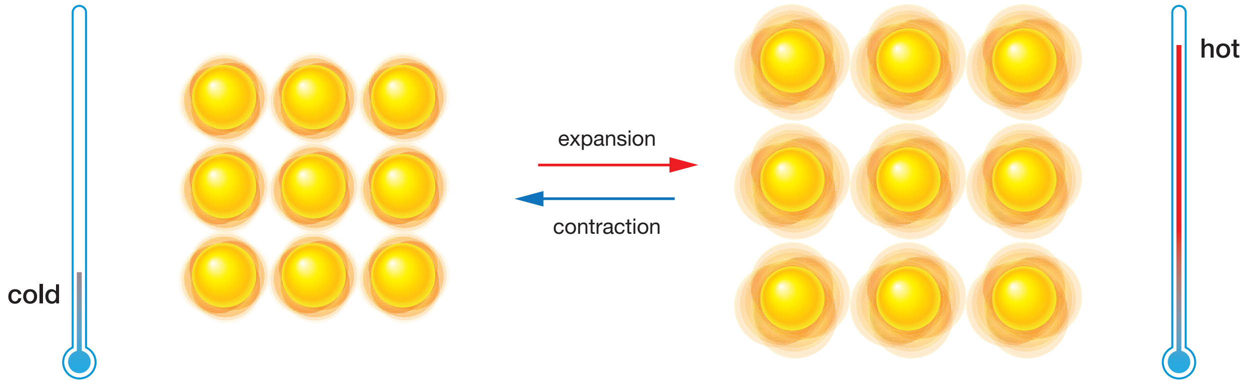


Figure 6.2.2 Increasing temperature causes the particles in solids and liquids to vibrate more. As are result, solids and liquids expand with increasing temperature.

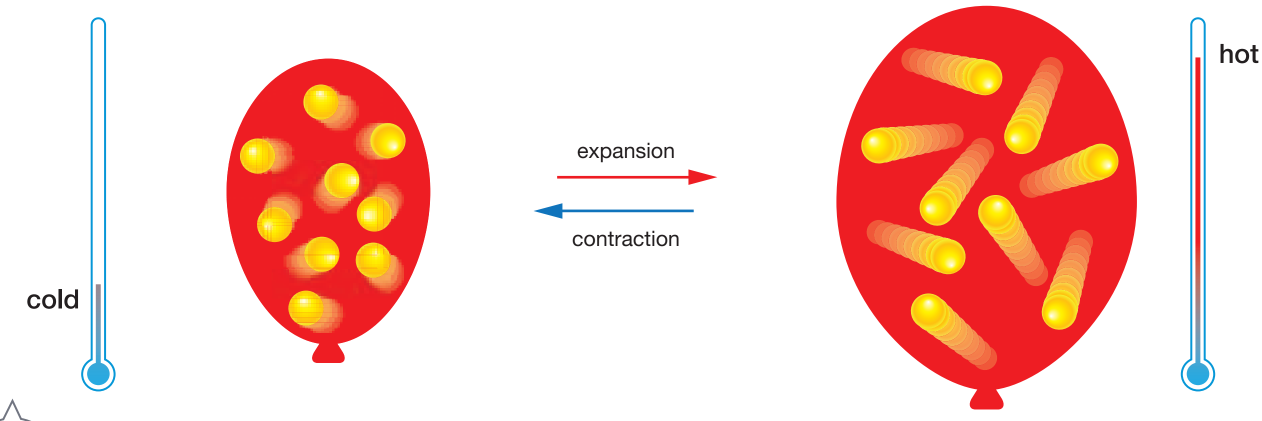


Figure 6.2.3 When the temperature of a gas is increased, the particles travel faster, hitting the sides of their container more frequently and with more force. If the container is flexible, such as a balloon, it will expand.

Changes of state and the particle model

The particle model explains changes of state by looking at the relationship between how the particles move and the attraction between them (Figure 6.2.4).

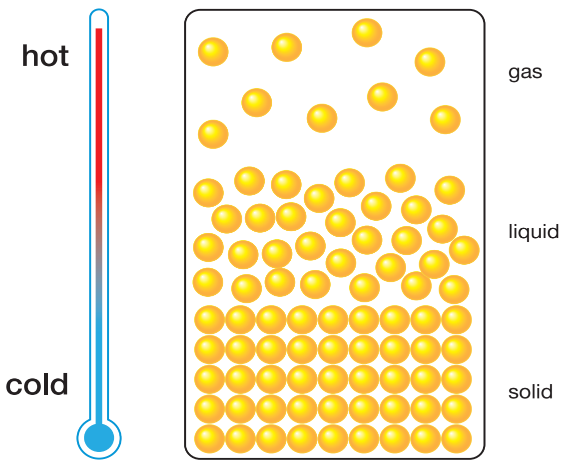


Figure 6.2.4 Increasing the temperature of a solid gives its particles enough energy to overcome the attraction between them and they form a liquid. Increasing the temperature of a liquid gives the particles enough energy to overcome the forces of attraction holding them together and they form a gas.

Melting and freezing

In a solid, the particles vibrate but are held in position by the forces of attraction between them. As the temperature increases, the vibrations increase and the solid expands. As the temperature is increased further, the solid melts. This is because the vibrations become so energetic that the attraction between the particles can no longer hold them in fixed positions. At this point, the particles become unstuck and start moving freely. However, there is still a small amount of attraction between the particles that holds them together as a liquid.

Freezing is the exact opposite of melting. As the free-moving particles in the liquid are cooled, the particles become slow and less energetic until the attraction between the particles is able to fix them in position, forming a solid.

Evaporation and condensation

Evaporation occurs when particles in a liquid escape from the surface of the liquid to form a gas. The particles in a liquid are stuck together by only weak forces of attraction. As the liquid particles are heated, they move faster until they are able to escape from the surface of the liquid to form a gas. In other words, the particles evaporate. At the boiling point, the particles within the liquid are moving so fast that they fly apart and form bubbles within the liquid.

Condensation occurs in reverse. As a gas is cooled, the particles move slower until a point where the forces of attraction between the particles can hold and stick them together to form liquid droplets.

2 p219

Mixing

Mixtures are formed when two or more pure substances are mixed together. A mixture of sand and iron filings is an example of a large-scale mixture that you can see. However, substances can also mix on the particle level to form solutions and alloys. This process of mixing can also be understood through the particle model.

Solutions and the particle model

Solutions are an example of a mixture. When a solute dissolves, the solvent particles surround the solute particles and carry them away as in Figure 6.2.5. As a result, the solute seems to disappear. However, the solute particles are not destroyed. They are simply spread so thinly throughout the solvent that they appear invisible. The individual solute particles are so small that they cannot be seen with even the most powerful optical microscope.

3 p219

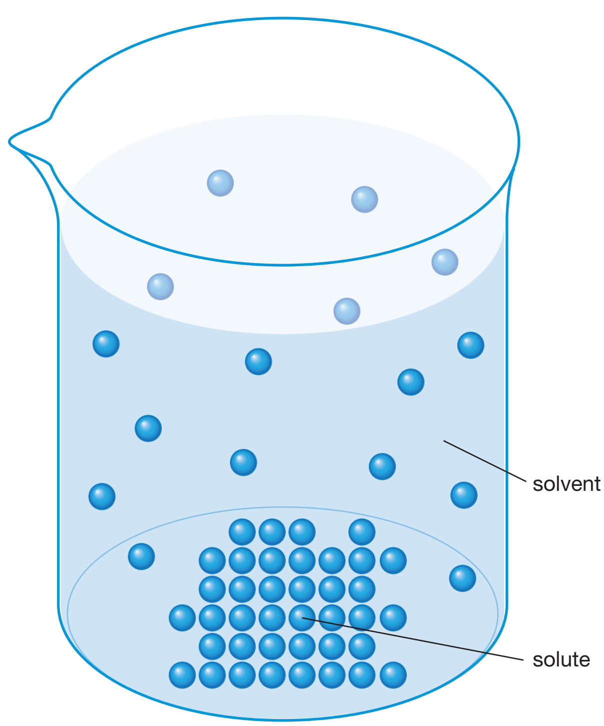


Figure 6.2.5 When a solute dissolves, its particles are spread evenly through the solvent. The particles are so small that they are invisible—even with the most powerful optical microscope.

Diffusion and the particle model

The particles of two gases or two liquids will mix evenly (or diffuse) without stirring. This process is known as **diffusion**. For example, when a bottle of perfume is opened on one side of a room, the perfume particles will diffuse through the air particles. That is why you will smell the perfume on the other side of the room after a short while.

During diffusion, gas particles travel in a zig-zag fashion. This is shown in Figure 6.2.6. The particles move in straight lines until they collide with another particle and then change direction. However, eventually the particles will be distributed evenly throughout the container. Increasing the temperature speeds up diffusion because the particles will travel faster.

Diffusion in liquids occurs less rapidly because the liquid particles are packed closer together. However, the liquids will ultimately spread evenly through one another.

4 p220

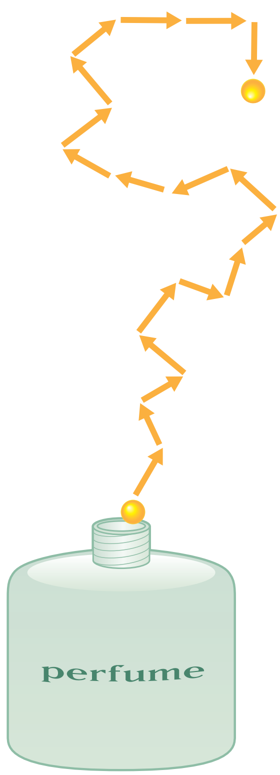


Figure 6.2.6 The perfume particle makes its way across the room in a zig-zag fashion, travelling in straight lines but changing direction whenever it collides with a particle in air.

6.3 6.4 6.5

INQUIRY science 4 fun

What smells?

Yes

How does a smell spread throughout a room?

Collect this …

- air freshener, perfume or aftershave in a spray bottle

Do this …

1. Everyone is to spread out around the room.
2. One person should spray a ‘squirt’ of the air freshener, perfume or aftershave into one corner of the room.
3. Raise your hand when you are able to smell the fragrance.

Record this …

Describe what happened.

Explain why you think this happened.

SciFile

Deadly diffusion

The process of liquid diffusion helps sharks track their prey. The blood from an injured animal diffuses through the water and can be detected by sharks several kilometres away due to their extreme sense of smell.

6.2 Unit review

Remembering

1 **List** the assumptions made by the particle model.

2 **Outline** the changes that occur when a gas is heated in a container that does not have flexible walls.

3 **Name** the opposite process for the following physical processes.
 a expansion
 b melting
 c evaporation
 d dissolving

Understanding

4 **Describe** the arrangements of particles in solids, liquids and gases.

5 **Outline** how the particle model explains expansion and contraction.

6 **Describe** how gas particles exert a pressure on their container.

7 **Explain** why you can smell something that is burning on the other side of the room.

8 **Describe** what happens to the particles in a solid as it is heated up to its melting point.

9 **Explain** why a solid seems to disappear when dissolved into a liquid.

Applying

10 **Use** the particle model to **explain** why your bike tyres may seem deflated on a cold day.

11 **Use** the particle model to **explain** why powerlines may sag more on a hot day.

Analysing

12 **Compare** dissolving and diffusion.

13 **Compare** the properties of solids, liquids and gases by copying and completing the following table.

	Solid	Liquid	Gas
Shape	Fixed		
Compressibility	Very low		
Strength of attraction between particles		Weak	
Movement of particles		Medium	
Space between particles			A large amount

Evaluating

14 **Propose** reasons why a scented candle spreads its smell quickly through a room while it is burning but not when it's unlit.

15 **Propose** a reason why gas barbecue bottles should not be left in the sun.

Creating

16 **Construct** diagrams showing what happens to gas particles in a solid container at low and high temperatures. Indicate on the diagrams which has high pressure and which has low pressure.

17 **Create** a short story from the perspective of a particle in a solid that is being heated. **Describe** what happens to you as the solid changes to a liquid and then to a gas.

Inquiring

1 Investigate Brownian motion, who it is named after and how it helped in the development of the particle model.

2 Investigate the life of Robert Boyle and his contribution to the particle theory.

6.2 Practical activities

1 Expansion and heat

Purpose

To observe how heating a gas causes it to expand.

Materials

- glass bottle
- hot tap water
- balloon

Procedure

1. Fill the glass bottle with hot tap water.
2. Pour out the hot water and quickly stretch a balloon over the mouth of the bottle.
3. Let the bottle cool or place it in a tub of cool water and observe what happens.

Results

Construct a series of diagrams that shows what happened.

Discussion

1. a **State** whether this is a chemical or physical change.
   b **Justify** your answer.
2. **Use** the terms *expansion* and *contraction* to **describe** your observations.
3. **Describe** what is happening inside the balloon using the particle model.

2 Observing sublimation

Purpose

To observe the change of state from a solid to a gas.

Materials

- small pieces of dry ice
- 1 L beaker
- hot and cold water
- drops of detergent
- tray
- gloves
- tweezers
- stopwatch

SAFETY

Dry ice is extremely cold. Never let it contact your skin directly because it will burn. Wear gloves.

Procedure

1. Using gloves and tweezers, place a piece of dry ice on your bench top and observe it as it sublimes.
2. Use the tweezers to slide the dry ice across your bench top and observe its motion.
3. Fill the bottom of the 1 L beaker with cold water and add a small amount of detergent.
4. Place some dry ice in the beaker and record how much time it takes to fill the beaker with bubbles.
5. Repeat steps 3 and 4, using hot water instead of cold.

Results

Record the time required to create 1 L of carbon dioxide bubbles with hot and cold water.

Discussion

1. a **Describe** how the dry ice moves across your bench top when you push it.
   b **Explain** why it moves in this manner.
2. a **State** whether increasing the temperature increases or decreases the rate of sublimation.
   b Use your results to **justify** your answer.
3. In this experiment, a change occurred in which bubbles were produced, normally a sign of a chemical change. **Explain** how you know that this activity shows a physical change and not a chemical change.

3 Dissolving and temperature

Purpose

To determine how the temperature of a solution affects the amount of solute that can be dissolved.

Materials

- 250 mL beaker
- thermometer
- stirring rod
- spatula
- 500 g sugar
- hot plate
- electronic balance
- watch-glass

Dissolving and temperature continued

6.2 Practical activities

Dissolving and temperature continued

Procedure

1. Fill the beaker with 100 mL of water.
2. Use the thermometer to measure the temperature and record the temperature in the table below.
3. Use the spatula to place approximately 20 g of sugar on to the watch-glass.
4. Use an electronic balance to measure the exact mass and record this value in the table below.
5. Add the sugar to the beaker of water and stir until it is dissolved.
6. Repeat steps 3–5 until the sugar will no longer dissolve.
7. Use the hotplate to increase the temperature to approximately 40°C, using the thermometer to measure the exact temperature.
8. Continue to add sugar approximately 20 g at a time until no more will dissolve.
9. Repeat steps 7 and 8 for temperatures of 60°C and 80°C.

Results

Record all your results in a table like that shown below.

Temp. (°C)	Mass of sugar added each time (g)							Total mass (g)

Discussion

1. **State** the name of the solute and the solvent in this experiment.
2. **Describe** how increasing the temperature changes the amount of solute that is dissolved.
3. **Use** the particle model to **propose** why the solubility changes in this way.
4. **Propose** a way the experiment could be changed to make it more accurate.

4 Rate of diffusion

Purpose

To observe the rate of diffusion with temperature.

Materials

- shallow tray
- overhead transparency with a grid of 2 cm × 2 cm squares printed on it
- hot and cold water
- food colouring
- dropper
- stopwatch

Procedure

This prac should be performed in pairs or small groups.

1. Place the gridded transparency on the bottom of the shallow tray.
2. Fill the tray with cold water.
3. With one person measuring time on the stopwatch, place a drop of food colouring and estimate how many squares are coloured after 5, 10 and 20 seconds.
4. Repeat the experiment with hot water.

Results

Copy and compete the following table.

	Number of squares		
	After 5 s	After 10 s	After 20 s
Cold water			
Hot water			

Discussion

1. **State** the name of the process that allows the food colouring to move through the water and whether this is a physical or chemical change.
2. **Compare** and **contrast** your observations for cold and hot water.
3. **Explain** your results in terms of the particle model.

6.3 Density

Some materials like lead, gold, granite and steel are very heavy for their size. Other materials like foam rubber, cork, balsa wood and feathers are so light that huge piles don't weigh much at all. Density measures how much matter is packed into a space. Density determines how heavy a handful of a substance will be and whether it floats or sinks in water.

INQUIRY

science 4 fun

Salty lava lamp

Yes

Can you make your own lava lamp?

Collect this ...

- tall glass or 250 mL beaker
- cooking oil
- salt shaker
- food dye
- water

Do this ...

1. Pour water into the glass or beaker until it is one-third full.
2. Add a few drops of food dye.
3. Pour in an equal quantity of cooking oil and observe which layer is on top.
4. Sprinkle salt into the glass and carefully observe what happens to the grains of salt.

Record this ...

Describe what happened.

Explain why you think this happened.

Density: a physical property

The physical properties of a substance describe things such as:

- its appearance
- the temperatures at which it melts, freezes and boils
- its hardness (how easily it is to scratch the material).

Density is one of the physical properties of a substance. Density depends on how heavy a substance is, but it is not the same as its weight or mass. For example, 1 kilogram of lead always weighs exactly the same as 1 kilogram of polystyrene, which weighs as much as 1 kilogram of feathers.

Density instead measures how much matter is packed into a space (Figure 6.3.1).

Figure 6.3.1

Stones are much more dense than feathers. This makes a small pile of stones weigh a lot more than a big pile of feathers.

Comparing densities

Density is normally measured in grams per cubic centimetre (symbol g/cm^3). Different substances pack different masses into the same size cubes, as shown in Figure 6.3.2. The densities of different solids, liquids and gases can be seen in Tables 6.3.1, 6.3.2 and 6.3.3.

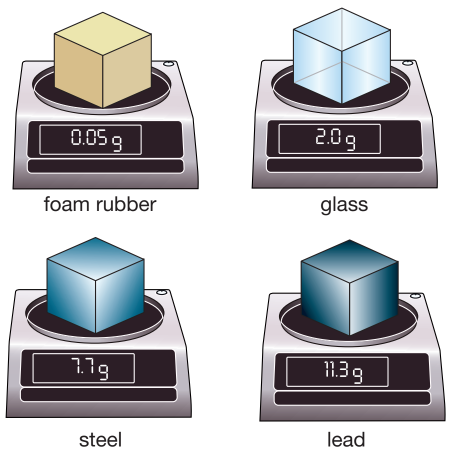


Figure 6.3.2 Density measures how many grams of matter are in a cubic centimetre of material. Density is measured in g/cm^3.

Table 6.3.1 Densities of gases

Gas	Density (g/cm^3)
Hydrogen (at 0°C)	0.000 09
Helium (at 0°C)	0.000 18
Air (at 40°C)	0.001 1
Air (at 0°C)	0.001 3
Oxygen (at 0°C)	0.001 4

Table 6.3.2 Densities of liquids at 25°C

Liquid	Density (g/cm^3)
Petrol	0.80
Vegetable oil	0.91
Water	1.00
Honey	1.36
Mercury	13.6

Table 6.3.3 Densities of solids

Solid	Density (g/cm^3)
Polystyrene foam	0.03
Wood (oak)	0.65
Concrete	2.40
Copper	8.90
Gold	18.9

Figure 6.3.3 The overall density of a hot-air balloon is less than the air around it, allowing it to rise into the sky.

Density and packing

The particles in a solid are usually packed a little more closely than they are in a liquid, making most solids a little denser than liquids of the same material.

The particles in a gas are very spread out, making gas the least dense of the three states. When a gas is heated, its particles spread out even further, lowering its density even more. For this reason, hot air is less dense than cold air and will rise above it (Figure 6.3.3). Likewise, bubbles of gas rise through liquid, as seen in Figure 6.3.4. This is why smoke rises—the air is hot and it carries soot and burnt material up with it.

Figure 6.3.4 Gases are much less dense than liquids and will bubble upwards through them.

Floating and sinking

Density determines how different substances will arrange themselves when mixed together. The densest substance will drop to the very bottom while the least dense will rise to the top.

Cooking oil, for example, floats on top of water because its density is less than that of water. A steel bolt will sink in water because its density is far greater than that of water.

The density of water is 1.0 g/cm^3. Anything more dense than this will sink when placed in water. Anything less dense will float on top of the water. Figure 6.3.5 shows two types of rock of very different densities.

Figure 6.3.5 Density determines whether things float or sink.

1 p228

Mass, volume and density

Density depends on the mass of the substance and the volume that mass takes up.

Mass

Mass measures how much matter is in a substance. It is sometimes incorrectly referred to as weight. Scientists use grams (g) to measure small masses such as a twig, a mouse or a spatula load of chemicals, kilograms (kg) for heavier masses such as a dog or a human, and tonnes (t) for even bigger masses such as a car or an aircraft. Mass is measured using a beam balance or electronic balance.

Volume

Volume is the amount of space a substance takes up. Several different methods can used to measure volume. The method used depends on what is being measured.

Regular solids

The volume of solids that have a regular shape (such as a cube, a box-shaped block, a sphere or a cylinder) can be calculated by measuring their dimensions (length, width, height or diameter), and then using the appropriate mathematical formula.

Volume is measured in cubic millimetres (mm^3) for really small objects such as crystals of salt, cubic centimetres (cm^3) for larger objects such as sugar cubes, and cubic metres (m^3) for much larger objects such as the amount of concrete in a building.

Calculating the volume of a rectangular prism

A rectangular prism is a box-shaped block, much like a shoe box. Its volume is calculated using the mathematical formula:

Volume = length × width × height

$V = lwh$

WORKED EXAMPLE

Calculating the volume of a rectangular prism

Problem

Calculate the volume of the shoe box shown in Figure 6.3.6.

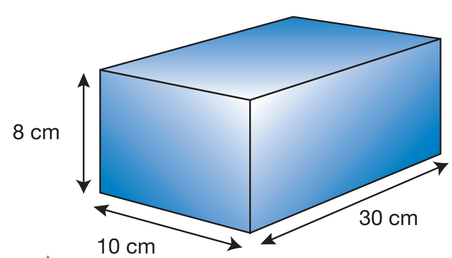


Figure 6.3.6 A rectangular prism

Solution

$V = lwh$

$= 30 \times 10 \times 8$

$= 2400\ cm^3$

Liquids

The volume of a liquid can be measured accurately using a measuring cylinder. The volume of liquids is usually measured in litres (L). For example, large bottles of soft drink come in 1.25 L and 2 L sizes. Smaller volumes are usually measured in millilitres (mL). For example, cans of soft drink normally hold a volume of 375 mL.

It is easy to convert between mL and cm^3 since 1 mL is exactly the same as 1 cm^3.

$$1 \text{ mL} = 1 \text{ cm}^3$$

Figure 6.3.7 Archimedes (about 287 to 212 BCE) was supposedly given the task of determining whether the wreath of his king was pure gold or not. To do this he needed to measure its volume and density without destroying it. He found the volume by submerging the wreath in water and measuring how much water was displaced.

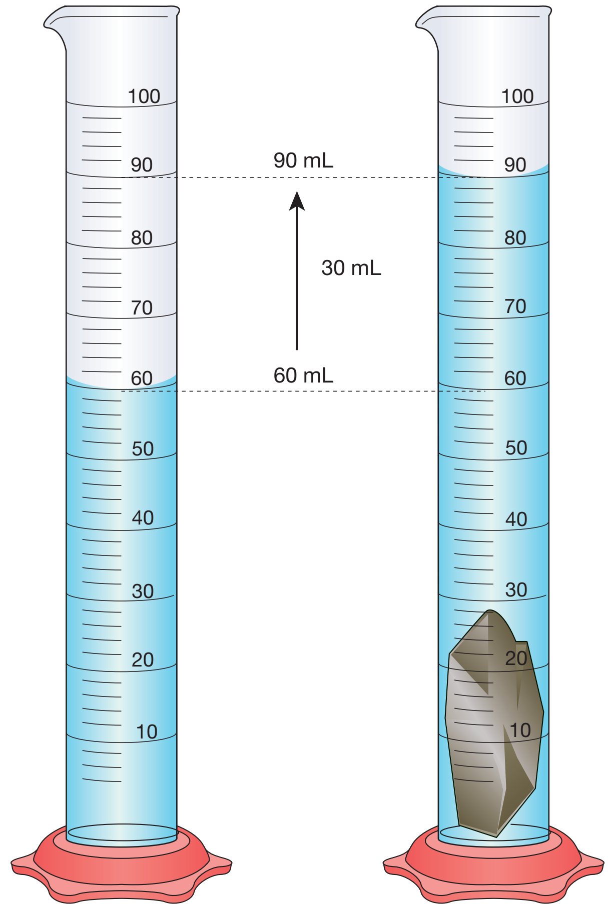


Figure 6.3.8 The volume of an irregular solid can be determined by dropping it into a measuring cylinder. If the level goes up by 30 mL then the solid must take up 30 mL.

SciFile

Danny Deckchair

Helium is a gas that is less dense than air. In 1982, US man Larry Walters tied 45 large helium-filled balloons to his aluminium garden chair and floated to a height of 4900 metres. After 45 minutes he punctured some of his balloons and crashed down among power lines. A similar flight was used in the plot of the 2003 Australian film *Danny Deckchair*.

Irregular solids

Many solids do not have predictable regular shapes. This means that you can't calculate their volume easily using a mathematical formula. Instead, their volume is measured by dropping them into water in a measuring cylinder. When you place a solid in water, it displaces (pushes upwards) a volume of water equivalent to its own volume. This is shown in Figure 6.3.8. That's why the level of water in the bath goes up when you get in.

Calculating density

Density is the amount of matter (or mass) packed into a certain space (volume). This can be represented by the mathematical formula:

$$\text{density} = \frac{\text{mass}}{\text{volume}}$$

Using symbols, this formula can be written as:

$$d = \frac{m}{V}$$

For density calculations, mass is normally measured in grams (g) and volume in cubic centimetres (cm^3). Using these units, density is measured in grams per cubic centimetre (g/cm^3).

6.6
2 p229

Using the density formula

The triangle in Figure 6.3.9 can be used to calculate the density, mass or volume of a substance.

Simply cover the quantity you want to calculate and do what the rest of the triangle indicates.

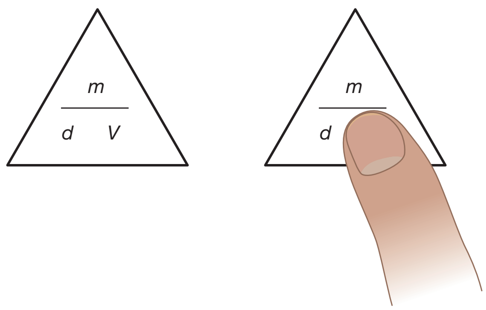


Figure 6.3.9 This triangle makes the density formula much easier to use. For example, to calculate volume, cover *V* with your finger. The part of the triangle not covered shows that you must divide mass by density.

Unit conversions

Scientists often need to convert one unit into another before they calculate the density of an object. In the SI or metric system, these conversions are carried out by multiplying or dividing by 10, 100, 1000 or some other factor of ten. Table 6.3.4 shows how to carry out some simple unit conversions.

Table 6.3.4 Converting mass and volume

Measurement	From	Convert by	To
Mass	tonne (t)	x 1000	kilogram (kg)
	kilogram (kg)	x 1000	gram (g)
Volume	litre (L)	x 1000	millilitre (mL)
	millilitre (mL)	x 1	cubic centimetres (cm^3)

WORKED EXAMPLE
Calculating density

Problem

The mass of the cube shown in Figure 6.3.10 is 72 g. Calculate its density and identify what the cube is most likely made from.

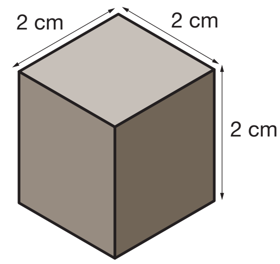
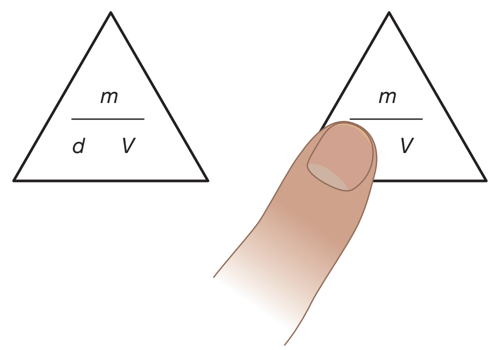


Figure 6.3.10 A 72 g cube of an unknown metal, with each side 2 cm long.

Solution

$$V = l \times w \times h$$
$$= 2 \times 2 \times 2$$
$$= 8\ cm^3$$

You want to calculate density so cover *d* in the triangle.



$$d = \frac{m}{V}$$
$$= \frac{72}{8}$$
$$= 9\ g/cm^3$$

This is roughly the density of pure copper and so the block was most likely copper.

Problem

Ice has a density of 0.9 g/cm^3. An ice cube has a mass of 9 g. Calculate the volume of the ice cube.

Solution

You need to calculate the volume, so cover the *V* in the density triangle.

$$V = \frac{m}{d}$$
$$= \frac{9}{0.9}$$
$$= 10\ cm^3$$

Icebergs

Water acts just like other liquids above 4°C and ice acts just like other solids below 0°C—it expands when heated and contracts when cooled. Between 0°C and 4°C, however, water does the exact opposite. Water at 4°C is denser than at any other temperature, and drops to the bottom of any pond, lake or swimming pool. Colder water floats on top of it, and any ice will float on the surface of that colder water. This is shown in Figure 6.3.11.

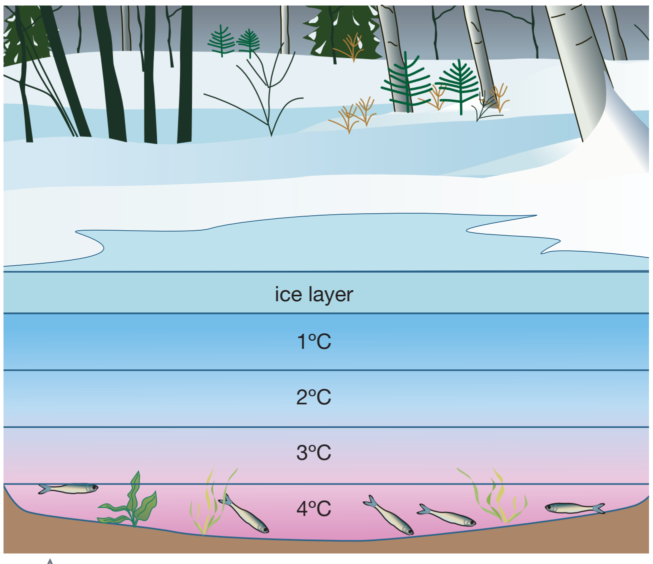


Figure 6.3.11 The water at the bottom of a pond, lake or swimming pool is always the densest. Even in freezing conditions, this layer is unlikely to drop below 4°C, giving fish some chance of survival.

Thicker, heavier chunks of ice will be partly submerged with only their top exposed, forming an iceberg. Despite their huge size and mass, icebergs float because their density is less than that of seawater, as shown in Table 6.3.5.

Table 6.3.5 Densities of water

Substance	Density (g/cm^3)
Ice	0.92 g/cm^3
Pure water	1.00 g/cm^3
Seawater	1.03 g/cm^3

Icebergs are incredibly dangerous to shipping because:

3 p230

- they shift with ocean currents and winds, often into shipping lanes.
- 80–90% lies hidden below the water as shown in Figure 6.3.12. This hidden part doesn't melt as fast as the ice above the water, and so it usually extends wider than the ice visible above the water line. Any ship that comes near can collide with this bulge, making a hole and sinking the ship.

Figure 6.3.12 A composite photo showing how much of an iceberg is hidden.

Although radar gives modern shipping some warning of a nearby iceberg, disaster can still happen. In 2007, the Canadian cruise ship MS *Explorer* (shown in Figure 6.3.13) hit submerged ice off Antarctica and sank.

Figure 6.3.13 When MS *Explorer* sank, all 154 passengers and crew were saved by a nearby cruise ship.

There was no radar in 1912 when RMS *Titanic* hit an iceberg and sank on its maiden (first) voyage from England to the USA. Although the crew knew that icebergs lay in its path and were watching out for them, it was a moonless night and the iceberg was only seen at the last moment.

6.3 Unit review

Remembering

1 **State** whether each of the following is true or false.
  a 1 kg of lead weighs more than 1 kg of polystyrene.
  b 1 kg of lead takes up less space than 1 kg of polystyrene.
  c 1 kg of lead has the same density as 1 kg of polystyrene.

2 **Recall** the densities of different states and forms of water by arranging the following in order from least dense to most dense:

  ice, seawater, pure water at 1°C, pure water at 4°C, water vapour

Understanding

3 **Describe** how mass is normally measured.

4 **Describe** how you could measure the volume of a:
  a cube of stone
  b rough pebble.

5 **Explain** why it is easier to float in the sea than in a freshwater lake.

Applying

6 **Use** the densities in Tables 6.3.1, 6.3.2 and 6.3.3 on page 222 to **explain** why:
  a lifesaving rings are usually made of polystyrene
  b helium balloons rise
  c firefighters recommend you do not run out of a burning house but that you drop to the floor and crawl out to escape
  d spilt petrol forms a slick on top of water.

7 **Calculate** the density of a solid that has a:
  a volume of 2 cm^3 and a mass of 3 g
  b mass of 8.4 g and a volume of 4 cm^3.

8 **Calculate** the volume of a solid that has a:
  a mass of 6 g and a density of 3 g/cm^3
  b density of 3 g/cm^3 and a mass of 39 kg (be careful!).

9 For each of the objects shown in Figure 6.3.14, **calculate** the:
  i volume (in cm^3)
  ii density (in g/cm^3).

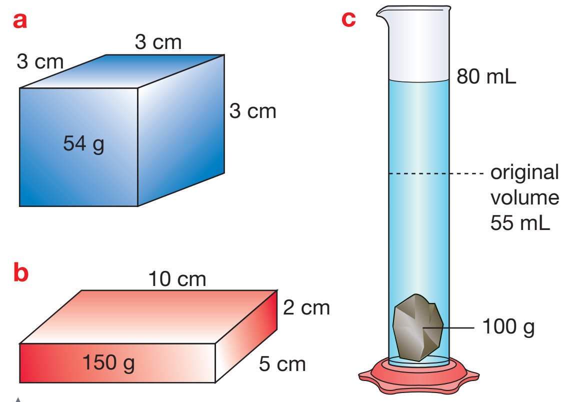


Figure 6.3.14

Evaluating

10 a **Predict** whether you are more or less dense than water.
  b **Justify** your answer.

11 Newspaper usually floats on water but then sinks after a while. **Propose** reasons why.

Creating

12 a **Construct** a diagram showing a measuring cylinder containing all the liquids shown in Table 6.3.2 on page 222.
  b Small lumps of wood, copper and gold were then dropped into the measuring cylinder. On your diagram, show where each would settle.

Inquiring

1 Find evacuation procedures recommended by your state fire service. Construct a simple poster that displays one evacuation procedure.

2 Construct a timeline that shows the important events in the sinking of RMS *Titanic*.

3 Search the internet for videos of icebergs calving (breaking off).

6.3 Practical activities

1 Density tower

Purpose

To construct a tower of different liquids, layered according to their densities.

Materials

- large measuring cylinder
- corn syrup or honey
- food dyes
- water
- vegetable oil
- ethanol or methylated spirits
- dishwashing liquid (coloured)
- a variety of small solids (such as a cornflake, single penne pasta, cork, sultana, bolt, small rubber stoppers, grape, Lego® block)
- digital camera or mobile phone with camera function (optional)

Procedure

1 Carefully squeeze the honey into the measuring cylinder so that it forms a layer at least 1 cm thick on the bottom.

2 Carefully squeeze or pour a similar quantity of dishwashing liquid into the measuring cylinder. Do this by tilting the cylinder and slowly pouring the dishwashing liquid down its side.

3 Choose a food dye that is a different colour from the dishwashing liquid and add a few drops of it to a small beaker of water. Tilt the cylinder again, and carefully pour the coloured water in to form a layer of about the same thickness as the others.

4 Use the same method to carefully pour a layer of vegetable oil on top of the coloured water.

5 Add a few drops of food dye to a small beaker of ethanol or methylated spirits. As before, make another layer by gently pouring the ethanol or methylated spirits down the side of the measuring cylinder.

7 Stand the measuring cylinder upright and allow the contents to settle.

8 Construct a labelled sketch showing the layering of liquids in the tower. Alternatively, photograph the tower with a digital camera or the camera function of your mobile phone.

9 Gently lower the small solids, one by one, into the measuring cylinder. Record on your diagram the level at which each one settles.

Discussion

1 **Explain** why:
   a the liquids formed layers
   b some objects floated and others sank.

2 **Identify** the least dense:
   a liquid
   b object.

3 **Explain** how you came to this conclusion.

4 **List** in order from most to least dense all the:
   a liquids you tested
   b solids you tested
   c liquids and solids you tested.

2 Density of an irregular shape

Materials

- 100 mL measuring cylinder
- a selection of irregularly shaped objects (such as a rubber stopper, bolt, pebble, pencil sharpener, piece of bark)
- short length of thin wire
- access to an electronic balance

Procedure

1. Make sure that every object that you are about to test fits inside the 100 mL measuring cylinder.
2. Use the electronic balance to measure the mass of each object. Record each mass in a table like that shown below.
3. Approximately half-fill the measuring cylinder with water as shown in Figure 6.3.15.
4. Record the start volume (in mL) in the measuring cylinder in your table. Make sure you read the water level from the bottom of the meniscus.
5. Carefully slide an object into the measuring cylinder. If it floats, use the piece of thin wire to push the object completely under the water.
6. Record the final volume (in mL) in the measuring cylinder.

Results

1. Copy the table below into your workbook. Use it to record your measurements and calculations.
2. Calculate the volume of the object (in mL) by subtracting the starting volume of water from the final volume of water. Convert your answer into cm^3 (remember 1 mL = 1 cm^3).
3. Use the formula $d = \frac{m}{V}$ to calculate the density (in g/cm^3) of each object you tested.

Discussion

List the objects from least dense to most dense.

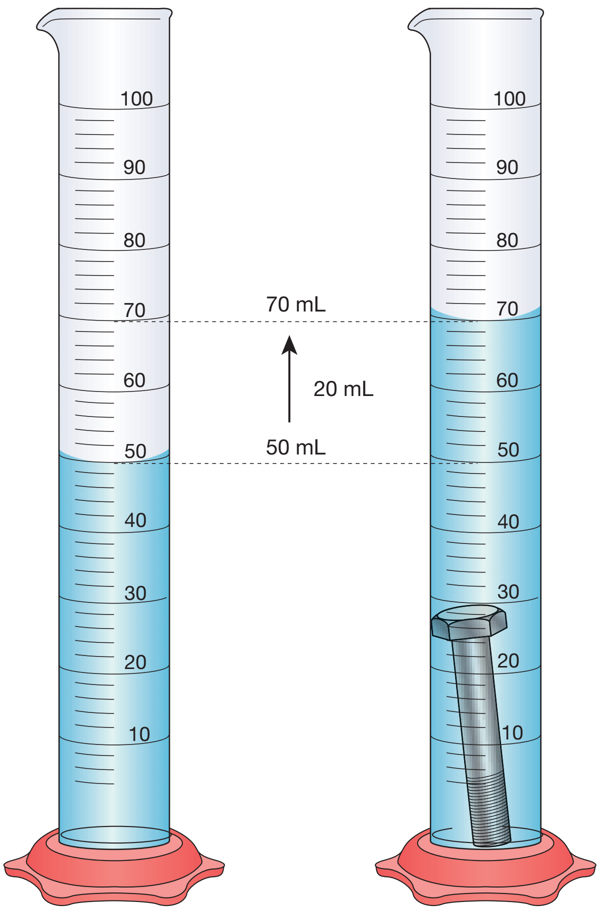


Figure 6.3.15

Column 1	Column 2	Column 3	Column 4	Column 5	Column 6	Column 7
Object	**Mass (g)**	**Initial volume (mL)**	**Final volume (mL)**	**Volume = column 4 – column 3 (mL)**	**Volume = column 5 (cm^3)**	**Density = column 2 ÷ column 6**

3 Icebergs and eggbergs

Purpose

To demonstrate how the different densities of fresh and salt water change how something floats.

Materials

For both parts

- water
- salt
- spoon

For Part A: Icebergs

- glass or 250 mL beaker
- 2 ice cubes
- plastic 30 cm ruler

For Part B: Eggbergs

- 2 tall glasses, tall beakers or large measuring cylinders, wide enough to take an egg
- fresh uncooked egg
- large spoon
- salt
- water

Method

Part A: Icebergs

1. Three-quarters fill the glass or beaker with water.
2. Slide in a cube of ice.
3. Use the ruler to measure how much of the ice cube is above water and how much is below water.
4. Use a calculator to calculate the percentage of the ice cube that lies below the water.

$$\% = \frac{\text{how much ice cube is below water}}{\text{total height of ice cube}} \times 100$$

5. Use the spoon to fish out the ice cube. Add salt to the water and stir. Keep stirring and adding salt until no more will dissolve.
6. Slide in another cube of ice and repeat your measurements.

Part B: Eggbergs

7. Fill two tall glasses, tall beakers or large measuring cylinders with approximately the same amount of water.
8. Add spoonfuls of salt to one glass until no more will dissolve.
9. Use the spoon to lower the fresh egg gently into the glass of fresh (non-salty) water. Observe what happens.
10. Lower the egg into the glass of salt water. Observe what happens.
11. Remove the egg and pour half the salt water out.
12. Very slowly add half the fresh water, making sure that it does not mix with the salt water. The best way of doing this is by slowly pouring the fresh water down the inside of the glass.
13. Use the spoon to lower the egg gently into the water. Watch where the egg settles.

Discussion

1. **List** the evidence that suggests that salt water is more dense than fresh water.
2. **List** the following substances in order from most dense to least dense:

   egg, fresh water, ice cube, salt water
3. **Use** the evidence found in this experiment to **explain** why icebergs are so dangerous to shipping.
4. Although many attempts were made to find the wreck of RMS *Titanic*, it was only discovered in 1985, 21 km away from its last recorded position. Some scientists think that the wreck may not have hit the bottom immediately but settled somewhere in the water above it. **Use** observations made in this activity to **explain** how this might have happened.

6.4 Chemical reactions

Chemical reactions play a vital role in your everyday life. When you take a deep breath, you start a series of chemical reactions that keep you alive. When you place a log on a fire, the chemical reactions keep you warm. Chemical reactions in a car's engine keep it running. Chemical reactions are also used to create new and useful materials such as fuels, plastics, foods and medicines.

Understanding chemical change

Matter is made up of particles. The very simplest of these particles are called **atoms**. Atoms are the building blocks from which all substances are made. Atoms can be thought of as hard spheres. Atoms rarely exist by themselves but usually group together to form clusters called **molecules**. A typical atom and molecule are shown in Figure 6.4.1.

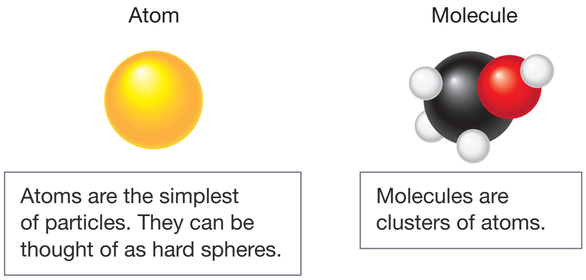


Figure 6.4.1 Chemical changes can be understood by assuming that the particles making up a substance are molecules made up of atoms.

This is known as the atomic theory of matter and was first written down by English chemist John Dalton in 1803. The five principles of Dalton's atomic theory of matter are as follows:

1. All matter is made up of hard, tiny, invisible particles called atoms.
2. Substances made of one type of atom are called elements. For example, the element carbon is made of carbon atoms.
3. The atoms of different elements can be distinguished by their different masses.
4. Atoms can combine to form new substances called molecules.
5. Atoms are not created or destroyed during chemical change.

From these basic principles, scientists know that the molecules in a pure substance all have the same number of atoms and are identical in size and shape. When a chemical change occurs, the atoms in the molecules rearrange to form a different combination of atoms. This is how new substances are produced during a chemical change.

For example, when carbon (C) is burnt in oxygen gas (O_2), an atom of carbon will combine with a molecule of oxygen and rearrange to form carbon dioxide gas (CO_2). This is shown in Figure 6.4.2.

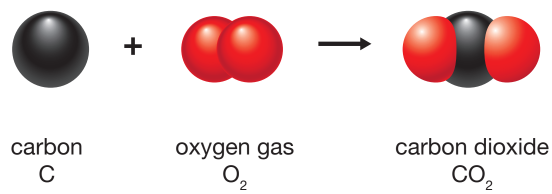


Figure 6.4.2 When carbon is burnt in oxygen gas, the atoms rearrange to form carbon dioxide.

If the atoms rearrange to form a new substance with a different colour, then a permanent colour change will be observed. If the atoms rearrange to form a new substance that is a gas, then you will see bubbles or smoke, or smell an odour. If the atoms rearrange to form a new substance that is a solid, then a precipitate might be observed.

During a chemical change, atoms are never created or destroyed—they simply rearrange. When hydrogen gas (H_2) burns in oxygen gas (O_2), the atoms in the hydrogen and oxygen molecules rearrange to form water (H_2O). However, if just one hydrogen molecule reacted with just one oxygen molecule, there would be an oxygen atom left over. Therefore, two molecules of hydrogen must combine with each oxygen molecule. Their atoms then rearrange to form two molecules of water as shown in Figure 6.4.3.

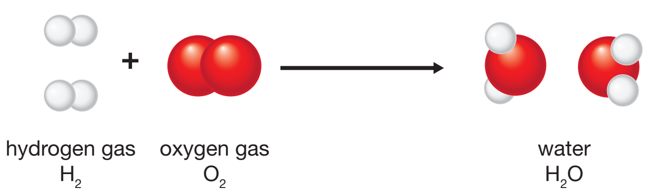


Figure 6.4.3 When hydrogen burns in oxygen, two hydrogen molecules react with one oxygen molecule to form two water molecules.

Chemical change and reactions

When a chemical change is observed, scientists say that a chemical reaction has occurred. A **chemical reaction** is when atoms rearrange to form new substances. The starting substances of a chemical reaction are known as the **reactants**. The substances that are formed are known as the **products**.

Atoms in chemical reactions

During a chemical reaction, the atoms in the reactants rearrange to form the products. The atoms can only rearrange—they cannot appear, disappear or change type. For example, when nitric oxide (NO) reacts with ozone gas (O_3), the atoms rearrange to produce nitrogen dioxide (NO_2) and oxygen gas (O_2). This chemical reaction is shown in Figure 6.4.4.

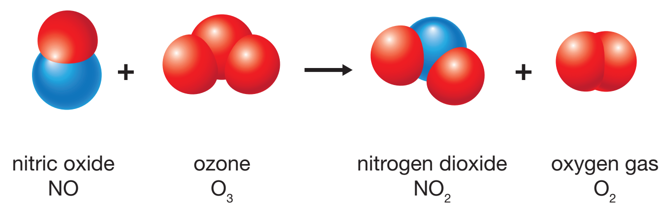


Figure 6.4.4 When nitric oxide is heated with ozone, atoms rearrange and combine to form nitrogen dioxide and oxygen gas.

In this chemical reaction, two reactants (nitric oxide and ozone gas) form two products (nitrogen dioxide and oxygen gas). In the reactants, there is one nitrogen atom and four oxygen atoms. The product molecules contain one nitrogen atom and four oxygen atoms. No new atoms were created and no atoms were destroyed. The chemical reaction simply rearranged the nitrogen and oxygen atoms.

Figure 6.4.5 shows what happens when sodium hydroxide (NaOH) reacts with hydrochloric acid (HCl). They rearrange to form sodium chloride (NaCl) and water (H_2O).

p241

Figure 6.4.6 shows that if magnesium metal (Mg) reacts with sulfuric acid (H_2SO_4), then their atoms rearrange to from two products—hydrogen gas (H_2) and magnesium sulfate ($MgSO_4$).

6.7

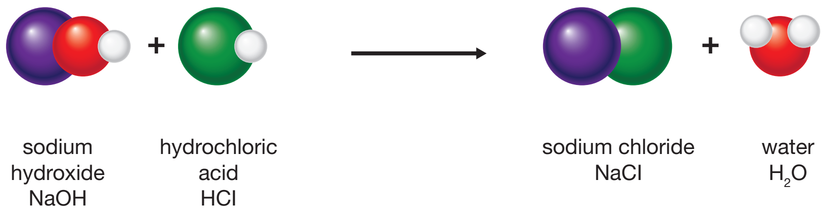


Figure 6.4.5

When sodium hydroxide is mixed with hydrochloric acid, the atoms rearrange to form sodium chloride and water.

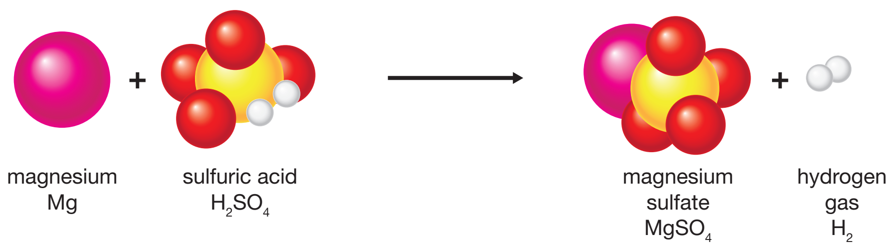


Figure 6.4.6

Reacting magnesium with sulfuric acid just rearranges atoms.

Energy in chemical reactions

Energy plays an important part in chemical reactions. The atoms in molecules and lattices are held together by chemical bonds that store energy. It is the energy in these bonds that is released when you use a battery, burn fossil fuels like petrol or convert food into energy for your body.

Spontaneous and non-spontaneous reactions

Some chemical reactions need a constant supply of energy while others will start and continue naturally. Reactions that proceed by themselves are known as spontaneous reactions. There are two types of spontaneous reactions. The first type can obtain enough energy from the surroundings to start and continue. Rusting is an example of this type of spontaneous reaction. Other spontaneous reactions need a kick-start from an external energy source but then produce enough energy to continue the reaction. Although you need an initial flame to start wood burning, the chemical reaction produces enough heat to continue the reaction spontaneously. The burning sparkler in Figure 6.4.7 is another example of a spontaneous reaction. It only stops when the fuel itself runs out.

Other chemical reactions require energy to be added constantly—otherwise the reaction will stop. These reactions are known as non-spontaneous reactions. A hydrogen fuel cell (like the one in Figure 6.4.8) converts water (H_2O) into hydrogen (H_2) and oxygen (O_2). Although this is a clean source of energy, it requires an electrical current to be applied constantly for the reaction to proceed.

Figure 6.4.7

A sparkler is an example of a spontaneous reaction. Once ignited, it continues spontaneously.

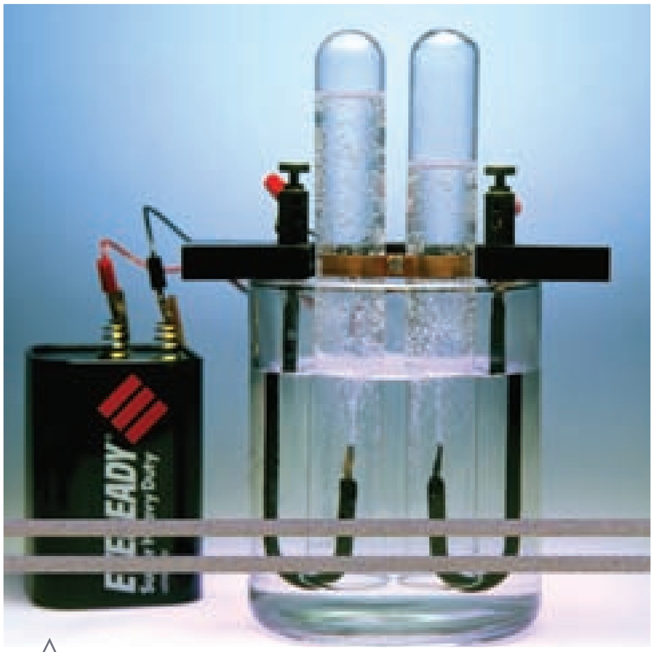


Figure 6.4.8

Water can be broken down into oxygen and hydrogen. This requires energy to be constantly put into the reaction. Therefore this reaction is non-spontaneous.

Representing chemical reactions

Describing chemical reactions in pictures or sentences is a very clumsy way of describing what is happening in a chemical reaction. For this reason, scientists have developed a shorthand way of showing what is happening. This shorthand version is called a chemical equation. Chemical equations make describing chemical reactions clearer, faster and easier. The simplest and most general chemical equation is:

$$\text{reactants} \longrightarrow \text{products}$$

The arrow ($\longrightarrow$) means that the reactants become the products. Chemical equations can be used to specify the reactants and products of any chemical reaction.

Word equations

The simplest way to write a chemical equation is to use the chemical names of the reactants and products.

For example, when hydrochloric acid (HCl) is mixed with sodium hydroxide (NaOH), they combine to form a solution of sodium chloride (NaCl) and water (H_2O). To describe this chemical reaction with a **word equation**, chemists write:

$$\text{hydrochloric acid} + \text{sodium hydroxide} \longrightarrow \text{sodium chloride} + \text{water}$$

WORKED EXAMPLE
Constructing word equations

Problem

When natural gas, methane (CH_4), is burnt in oxygen gas (O_2), it produces carbon dioxide (CO_2) and water (H_2O). Write a word equation for this reaction.

Solution

The reactants are methane and oxygen and the products are carbon dioxide and water. The general form for a chemical reaction is:

$$\text{reactants} \longrightarrow \text{products}$$

Replacing the reactants and products with their chemical names gives the word equation:

$$\text{methane} + \text{oxygen gas} \longrightarrow \text{carbon dioxide} + \text{water}$$

Formula equations

Even word equations are quite clumsy to write. For this reason chemists write **formula equations** instead. In a formula equation, the chemical name for a reactant or product is replaced with its chemical formula. In the case of hydrochloric acid (HCl) reacting with sodium hydroxide (NaOH) to form sodium chloride (NaCl) and water (H_2O), the formula equation is simply:

6.8

$$HCl + NaOH \longrightarrow NaCl + H_2O$$

WORKED EXAMPLE
Constructing formula equations

Problem 1

When a sodium chloride (NaCl) solution is mixed with a silver nitrate ($AgNO_3$) solution, white silver chloride (AgCl) is produced, leaving behind a sodium nitrate ($NaNO_3$) solution. Write the word and formula equations for this chemical reaction.

Solution

Reactants: sodium chloride (NaCl), silver nitrate ($AgNO_3$)

Products: silver chloride (AgCl), sodium nitrate ($NaNO_3$)

The general equation for a chemical reaction is:

$$\text{reactants} \longrightarrow \text{products}$$

This makes the word equation:

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

Replace the chemical names with the chemical formulas to get the formula equation:

$$NaCl + AgNO_3 \longrightarrow AgCl + NaNO_3$$

Problem 2

When sulfuric acid (H_2SO_4) is added to sodium carbonate (Na_2CO_3), it produces sodium sulfate (Na_2SO_4), carbon dioxide (CO_2) and water (H_2O). Write the word and formula equations for this chemical reaction.

Solution

Reactants: sulfuric acid (H_2SO_4), sodium carbonate (Na_2CO_3)

Products: sodium sulfate (Na_2SO_4), carbon dioxide (CO_2), water (H_2O)

The general equation for a chemical reaction is:

$$\text{reactants} \longrightarrow \text{products}$$

This makes the word equation:

$$\text{sulfuric acid} + \text{sodium carbonate} \longrightarrow \text{sodium sulfate} + \text{carbon dioxide} + \text{water}$$

Replace the chemical names with the chemical formulas to get the formula equation:

$$H_2SO_4 + Na_2CO_3 \longrightarrow Na_2SO_4 + CO_2 + H_2O$$

Chemistry in context: corrosion

Corrosion is a chemical reaction that costs the world billions of dollars every year. The term *corrosion* refers to chemical reactions where a metal reacts with oxygen. The most common and most costly form of corrosion is rusting.

The properties of metals determine how they are used. For example, iron is strong, making it ideal for ships and bridges. In contrast, aluminium is light and so is used in aircraft and drink cans.

The chemical properties of some metals make them more likely to corrode.

Rusting

Rusting is the name given to the corrosion of iron. In this reaction, iron metal reacts with oxygen in the air or water to produce the orange-brown compound iron(III) oxide (Fe_2O_3), which is more commonly known as rust. The word equation for this reaction is therefore:

$$\text{iron} + \text{oxygen} \longrightarrow \text{iron(III) oxide}$$

This chemical reaction can destroy anything made from iron such as cars, ships and bridges.

The true cost of iron

Rusting is an example of an unwanted and very destructive chemical reaction. Although iron is cheap to make, the CSIRO estimates that corrosion costs Australia $13 billion each year.

SciFile

Rust prevention

The impact of rusting on the world's iron structures is so extensive and costly that scientists have put a lot of effort into understanding how rusting occurs and how it can be prevented. The rusting of iron is usually minimised in three ways:

2 p241

- coating
- galvanising
- alloying.

Coating

The simplest and cheapest way to protect iron from rusting is to coat the iron with a material such as paint, plastic or another metal such as chromium. The coating forms a barrier between the iron metal and the oxygen in the atmosphere. As a result, rusting cannot occur. This is a cheap method of protection that is useful for large structures such as the Golden Gate Bridge (Figure 6.4.9) or large mass-produced items such as cars. However, if the coating is scratched or cracked, then the barrier is removed and rusting will occur once more.

Figure 6.4.9 Despite its name, the Golden Gate Bridge in San Francisco is actually painted red. It takes a team of 38 painters to maintain the paint coating and prevent corrosion.

Galvanising

If the iron is coated with zinc metal, then the zinc coating will continue to protect the iron even if the coating is scratched or cracked. This form of protection is known as galvanisation. Galvanisation works because oxygen reacts more easily with zinc than with iron. Therefore, even when the zinc is scratched and the iron is exposed to oxygen, the oxygen will not react with the iron but will react with the zinc instead. This form of protection is useful for some building materials such as nails, pipes or corrugated iron sheets as shown in Figure 6.4.10.

Figure 6.4.10 Galvanised iron is useful for producing building materials such as sheets of corrugated iron.

Alloying

If iron is mixed with other metals or carbon, then its chemical and physical properties can change dramatically. These mixtures of metals are known as alloys. Iron is almost never used in its pure form because pure iron is too soft and undergoes corrosion rapidly. For this reason it is usually mixed with carbon and other metals to produce steel, which is harder and more resistant to corrosion. Stainless steel is an alloy of iron that is particularly rust-resistant. It is a mixture of about 90% iron and 10% chromium. Stainless steel is more expensive than other alloys of iron so is mostly used for small items such as cutlery, watches and pots like those in Figure 6.4.11. However, it has been used in the construction of statues and bridges and to add decoration to buildings.

Figure 6.4.11 Stainless steel is an alloy of iron and chromium. It is much more expensive than other alloys of steel so is mainly used for small household items.

Corrosion of other metals

Aluminium (Al) also reacts easily with oxygen but does not corrode in the same way as iron. When aluminium reacts with oxygen, it forms a thin layer of aluminium oxide (Al_2O_3) on the surface. Unlike rust, the aluminium oxide coating forms a protective layer that prevents oxygen from reacting with the aluminium metal underneath as shown in Figure 6.4.12. As a result, aluminium can be used as a building material, without the need to apply any extra protection against corrosion.

Most metals will react with oxygen and therefore undergo corrosion. Some metals such as sodium (Na) and potassium (K) react with oxygen very rapidly and explosively to form sodium oxide (Na_2O) and potassium oxide (K_2O). Other metals such as zinc (Zn), copper (Cu) and silver (Ag) corrode more slowly with oxygen to produce zinc oxide (ZnO), copper(II) oxide (CuO) and silver oxide (Ag_2O). Metals such as gold (Au), platinum (Pt) and mercury (Hg) don't react with oxygen at all and are therefore referred to as noble metals.

INQUIRY science 4 fun

A cheap silver oxide remover

No

Can you remove the silver oxide from a necklace with simple household ingredients?

Collect this ...

- tarnished silver necklace
- tray
- aluminium foil
- bicarbonate of soda
- water
- electric kettle
- tongs

Do this ...

1. Place the aluminium foil in the tray.
2. Place the tarnished necklace on the aluminium foil and cover it with bicarbonate of soda.
3. Boil water in the electric kettle.
4. Pour the boiling water over the necklace.
5. Remove the necklace with tongs and observe.

Record this ...

Describe what happened.

Explain why you think this happened.

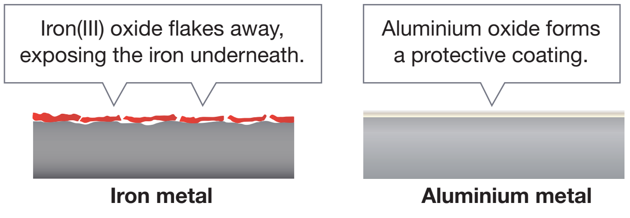


Figure 6.4.12 When iron corrodes, the oxide breaks away from the metal, allowing more oxygen in to attack the metal beneath. When aluminium corrodes, the oxide forms a perfect coating that protects the metal underneath from further corrosion.

SCIENCE AS A HUMAN ENDEAVOUR

Nature and development of science

The alchemists

Figure 6.4.13 The alchemists studied chemical and physical change but did not use scientific methods. This distinguishes them from modern-day chemists.

Today, the study of physical and chemical changes is called **chemistry** and the people who study them are known as chemists. Before chemistry, the study of physical and chemical changes was surrounded by myth and magic. These studies were performed by a group of people called alchemists. Today, scientists know that most of the things alchemists tried to achieve are technically impossible. However, many of the techniques they developed became the foundations of modern chemistry.

Who were the alchemists?

The study of alchemy can be traced back to ancient Greece and the Roman Empire. However, alchemy began to flourish in the Islamic world during the Middle Ages, and then in Europe from the 13th to the 18th centuries. Figure 6.4.13 shows what an alchemist's workshop might have looked like. Although alchemy was shrouded in mysticism, it attracted the attention of some of the brightest scientists of the time including Sir Isaac Newton (Figure 6.4.14) and Robert Boyle.

6.9

Figure 6.4.14 The famous scientist Sir Isaac Newton discovered important laws of forces, gravity and optics. He was also interested in alchemy.

ENDEAVOUR

In addition to the study of physical and chemical changes, alchemists aimed to achieve ultimate wisdom (knowing everything there is to know) and eternal life (the ability to live forever). It is this spiritual side of alchemy that distinguishes it from chemistry today. Another difference is that alchemists studied physical and chemical change by trial and error rather than using scientific method.

The word *alchemy* comes from the Arabic word *alkimia*, which means 'the art of transformation'. This is because alchemists tried to transform base metals such as lead into precious metals such as gold and silver. Alchemists also tried to discover a universal solvent (a liquid that could dissolve anything), the elixir of life for immortality and the panacea (a single medicine that could cure all disease).

Pee is for phosphorus

SciFile

In 1669, German alchemist Hennig Brand accidently discovered a new element while attempting to extract gold from urine. He collected 1100 litres of urine and let it rot until it produced a terrible smell. He then boiled it down into a paste and passed the vapours through water. Instead of gold, he obtained just 60 grams of a white substance that glowed in the dark. He'd discovered phosphorus, symbol P.

Figure 6.4.15 Phosphorus was discovered by an alchemist by accident in 1669.

The philosopher's stone

The philosopher's stone is one of the most well-known myths of the alchemists. It was a stone or tool that was believed to have the ability to change lead into gold. Others also believed that the philosopher's stone could give ultimate wisdom and eternal life. As a result, it was one of the most sought-after objects by alchemists in Europe. Today, it is regarded as myth but is often referred to in popular culture including the book and movie *Harry Potter and the Philosopher's Stone.*

Alchemy and modern chemistry

Through their studies, the alchemists developed many tools and techniques for studying and processing substances that are still in use today. An example is shown in Figure 6.4.16. Although the alchemists did not contribute a lot to our knowledge of chemistry, they did begin the first studies of matter and therefore can be considered the forefathers of modern chemistry.

Then in the 18th and 19th centuries, scientists began to realise that everything was in fact made of invisible particles called atoms and could explain all physical and chemical changes without the need for magic and mysticism. This marked the beginning of chemistry and the end for alchemy. As people's understanding of matter grew, it became increasingly apparent that the idea of transforming base metals was only a myth. So was eternal life.

Figure 6.4.16 Some laboratory processes and techniques developed by alchemists are still used today.

6.4 Unit review

Remembering

1 **State** where the chemical potential energy released from a reaction is stored.

2 **State** what the symbol ($\longrightarrow$) refers to in a chemical reaction.

3 **List** the five principles of the atomic theory of matter suggested by John Dalton.

4 **Recall** the word equation for the corrosion of iron.

5 **State** whether or not atoms can be created and destroyed during a chemical reaction.

Understanding

6 **Explain** what happens to the atoms during a chemical change.

7 **Explain** the difference between reactants and products.

8 **Explain** why iron structures need protection from corrosion while aluminium structures do not.

9 **Explain** why lighting a sparkler is considered a spontaneous reaction even though it doesn't start spontaneously.

10 **Explain** why scientists use chemical equations to describe chemical reactions.

11 **Explain** the difference betweem alchemy and chemistry.

12 **Describe** what is happening in Figure 6.4.17.

Applying

13 During the process of photosynthesis, plants use sunlight to convert carbon dioxide from the atmosphere and water into the oxygen we breathe and glucose in the foods we eat. **Use** a word equation to **summarise** this chemical reaction.

14 When iron oxide is heated in the presence of carbon monoxide, they combine to produce carbon dioxide and pure iron metal. **Use** a word equation to **summarise** this chemical reaction.

15 When magnesium sulfate solution ($MgSO_4$) is mixed with a solution of barium hydroxide ($Ba(OH)_2$), white barium sulfate ($BaSO_4$) is produced, leaving behind a solution of magnesium hydroxide ($Mg(OH)_2$). **Use** word and formula equations to **summarise** this chemical reaction.

16 **Identify** the word equations (a–e) with its formula equations (i–v) and match the equivalent reactions.

a sodium hydroxide + hydrochloric acid $\longrightarrow$ sodium chloride + water

b copper sulfate + barium chloride $\longrightarrow$ copper chloride + barium sulfate

c hydrogen + sulfur $\longrightarrow$ hydrogen sulfide

d nitric acid + lithium hydroxide $\longrightarrow$ lithium nitrate + water

e sulfuric acid + copper carbonate $\longrightarrow$ copper sulfate + carbon dioxide + water

i $H_2 + S \longrightarrow H_2S$

ii $CuSO_4 + BaCl_2 \longrightarrow CuCl_2 + BaSO_4$

iii $H_2SO_4 + CuCO_3 \longrightarrow CuSO_4 + CO_2 + H_2O$

iv $NaOH + HCl \longrightarrow NaCl + H_2O$

v $HNO_3 + LiOH \longrightarrow LiNO_3 + H_2O$

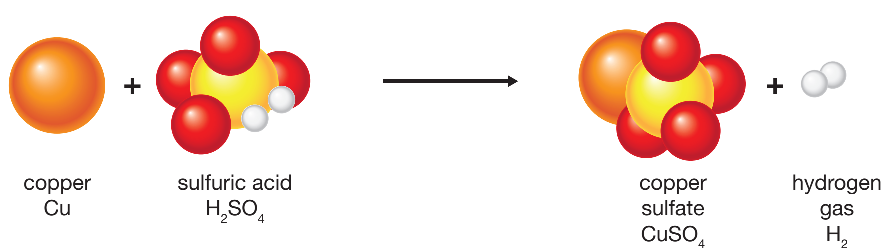


Figure 6.4.17

6.4 Unit review

Analysing

17 For each formula equation in Question 16:

- **a** **calculate** the number of each type of atom in the reactants
- **b** **calculate** the number of each type of atom in the products
- **c** **state** what you notice about the number of each type of atom on each side of the formula equation.

18 **Compare** spontaneous and non-spontaneous reactions by listing their similarities and differences.

Evaluating

19 **Propose** why alchemists are not considered chemists even though they worked with chemicals and used similar techniques.

12 You are an engineer asked to build a second bridge across Sydney Harbour.

- **a** **Propose** whether you should use painted iron, galvanised iron, stainless steel or aluminium.
- **b** **Justify** your answer.

Inquiring

1 Construct a labelled diagram of a blast furnace used to extract iron from its ore.

2 Research the process of respiration.

- **a** What is respiration?
- **b** Why is respiration important?
- **c** What are the reactants?
- **d** What are the products?
- **e** Where do you get the reactants?
- **f** What happens to the products?

Figure 6.4.18 Respiration in your body would not happen without the oxygen your lungs breathe in.

6.4 Practical activities

1 A precipitation reaction

Purpose
To observe the precipitation of barium sulfate.

Materials
- large test-tube
- barium hydroxide ($Ba(OH)_2$) solution
- magnesium sulfate ($MgSO_4$) solution
- 2 droppers

SAFETY
All chemicals should be treated as toxic.

Procedure
1. Using a dropper, add about 2 cm of barium hydroxide to the test-tube.
2. Using the second dropper, add about 2 cm of magnesium sulfate solution.

Results
Record your observations.

Discussion
1. **State** what indications there are that a chemical reaction is taking place.
2. Given that the precipitate is barium sulfate ($BaSO_4$), **predict** what other chemical may be produced during this reaction.
3. With this knowledge, **construct** the word and formula equations for this reaction.

2 Rust prevention

Purpose
To study the conditions under which rusting occurs.

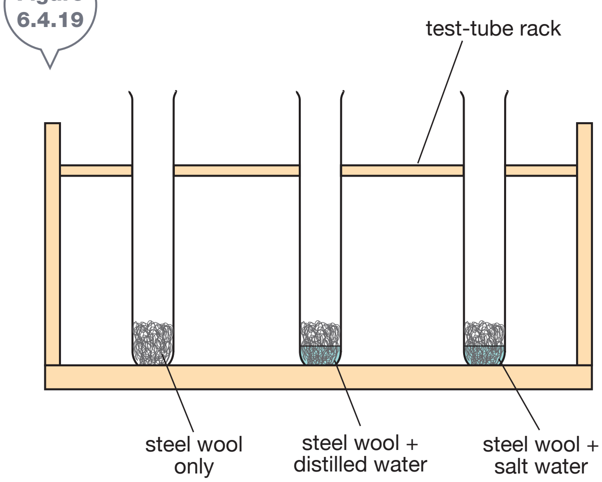
Materials
- 3 large test-tubes and test-tube rack
- 3 small pieces of steel wool
- distilled water
- sodium chloride solution (salt water)

Procedure
1. Place one piece of steel wool in each of the three test-tubes.
2. Do not add anything else to the first test-tube.
3. To the second test-tube, add distilled water so that the wool is half submerged.
4. To the third test-tube, add salt water solution so that the wool is half submerged (Figure 6.4.19).
5. Observe the reaction over the next 5 days.

Results
Copy and complete the following table.

Day	Observations		
	No water	Water	Water + salt

Figure 6.4.19



Discussion
1. **State** what evidence there is that a chemical reaction has taken place.
2. **Recall** the word and formula equations for rusting.
3. **List** what is necessary for rusting to occur.
4. **State** what else helps rusting to occur.
5. **Explain** why iron structures such as car bodies rust faster in coastal areas rather than inland.

6 Chapter review

Remembering

1 **List** four examples of physical changes and four of chemical changes.

2 **List** four things that indicate a chemical change has occurred.

3 **List** the changes of state that occur as you heat a solid to a gas and cool a gas to a solid.

4 **List** three ways that iron can be protected from corrosion.

Understanding

5 **Explain** why iron structures corrode while aluminium structures do not even though both metals react easily with oxygen.

6 **Outline** the physical changes that aluminium undergoes when it is recycled.

7 **Discuss** whether a kilogram of feathers has the same mass as a kilogram of rocks.

8 **Describe** how you would measure the volume of an irregular solid like a blob of clay.

Applying

9 **Use** the particle model to **describe** what happens to the particles of a solid as it is heated to form a liquid and then heated further to form a gas.

10 When methane gas is burnt in oxygen they produce carbon dioxide and water. **Use** a word equation to **describe** this reaction.

11 When magnesium metal (Mg) is combined with sulfuric acid (H_2SO_4) they react to form hydrogen gas (H_2) and magnesium sulfate ($MgSO_4$). **Use** word and formula equations to **describe** this reaction.

12 **Identify** which of the following formula equations match the word equations.

a $KOH + HCl \longrightarrow KCl + H_2O$
b $H_2SO_4 + CaCO_3 \longrightarrow CaSO_4 + CO_2 + H_2O$
c $C + O_2 \longrightarrow CO_2$
d $HCl + AgNO_3 \longrightarrow AgCl + HNO_3$
e $CuCl_2 + Pb(NO_3)_2 \longrightarrow Cu(NO_3)_2 + PbCl_2$

i carbon + oxygen $\longrightarrow$ carbon dioxide
ii potassium hydroxide + hydrochloric acid $\longrightarrow$ potassium chloride + water
iii sulfuric acid + calcium carbonate $\longrightarrow$ calcium sulfate + carbon dioxide + water
iv copper chloride + lead nitrate $\longrightarrow$ copper nitrate + lead chloride
v hydrochloric acid + silver nitrate $\longrightarrow$ silver chloride + nitric acid

13 A mobile phone has a mass of 80 g and a volume of 40 cm^3. **Calculate** the density of the mobile phone.

Analysing

14 **Classify** the following as physical or chemical changes: rotting apples, breaking a plate, dissolving copper sulfate, baking bread, making popcorn, melting butter, boiling oil, bending a paperclip, mixing paint, fermenting wine.

Evaluating

15 **Propose** why steel bridges are painted and not galvanised to protect them from corrosion.

Creating

16 **Use** the following ten key terms to **construct** a visual summary of the information presented in this chapter.

matter	mixing
physical change	chemical reaction
chemical change	atoms
change of form	molecules
change of state	chemical equation

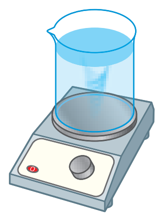
6.10

Thinking scientifically

Q1 During a physical change, no new substances are produced. Which of these processes is *not* an example of a physical change?

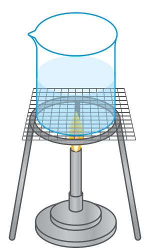
A

B



C

D



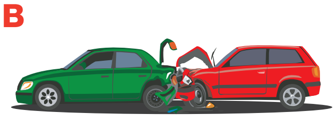
Q2 When a chemical change occurs new substances are produced. This may result in:

- a permanent colour change
- a gas being produced
- a solid precipitate being produced
- energy being produced or absorbed in the form of heat or light.

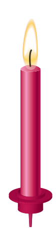
Identify which of these processes is *not* an example of a chemical change.

A

B



C



D

Q3 When Sanjay mixes sodium bicarbonate with vinegar, a chemical reaction occurs that releases a large amount of carbon dioxide gas. What would be the most obvious indication that a chemical change has occurred?

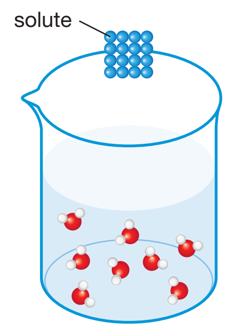
A a permanent colour change

B bubbles produced

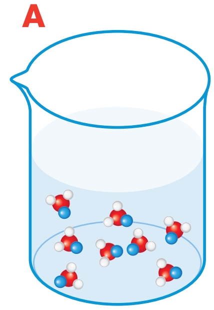
C precipitate formed

D heat produced

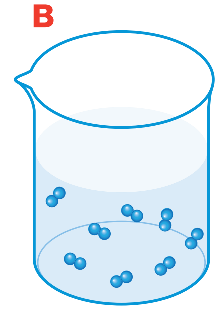
Q4 When a solute is dissolved in a solvent, the solute particles are distributed evenly through the solvent. Which of the following diagrams best represents what happens when the particles of the blue solute are dissolved in a beaker of water?



A



B



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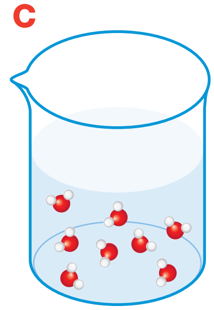
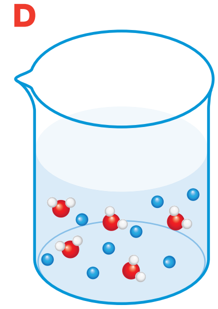


D



Glossary

Unit 6.1

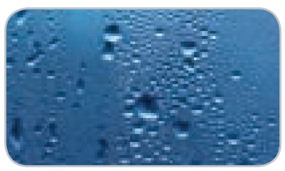
Boiling: a change of state where a liquid is heated and changes to a gas within the liquid

Boiling

Chemical change: a change that results in a new substance being formed

Condensation: a change of state where a gas is cooled and forms a liquid

Contraction: a decrease in size



Condensation

Deposition: a change of state from gas to solid

Endothermic: describes a process that absorbs energy from the surroundings

Evaporation: the change of state where a liquid changes to a gas at the surface of the liquid

Exothermic: describes a process that gives off energy in the form of heat, light or sound

Exothermic

Expansion: an increase in size

Freezing: the change of state from liquid to solid

Melting: the change of state when a solid is heated and forms a liquid

Physical change: a change that does not result in a new substance being produced

Precipitate: a solid formed during a chemical change

Precipitate

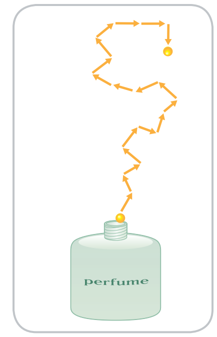
Solidification (freezing): the change of state from liquid to solid

Solution: a mixture where the substances are mixed on the atomic level

Sublimation: a change of state from solid to gas

Unit 6.2

Diffusion: a process where two liquids or two gases mix due to the motion of their particles



Diffusion

Particle model: a simplified representation of solids, liquids and gases

Unit 6.3

Density: a measure of the mass per unit volume of a substance, $d = \frac{m}{V}$ (unit: g/cm^3)

Displaces: pushes upwards, as in pushes water upwards

Mass: measures how much matter is in a substance (unit: g)

Volume: measures how much space is occupied by a substance (units: mL or cm^3)

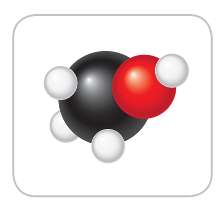
Unit 6.4

Atom: the smallest building block that makes up all substances

Chemical reaction: when atoms rearrange to form new substances

Formula equation: a type of chemical equation where the chemicals in the reaction are represented by their chemical formulas

Molecule: a cluster of atoms



Molecule

Products: the substances formed in chemical reactions

Reactants: the starting substances in chemical reactions

Word equation: a type of chemical equation where the chemicals are represented by the chemical names