CONTENTS

UNIT 1

6 Chemical analysis, assessing purity and separating mixtures

Specification points

This chapter covers specification points 1.9.1 to 1.9.18. It covers purity, separation of mixtures, potable water and identification of ions using chemical tests. You can find related work on writing ionic equations, which is useful for ion tests, in the appendix.

Assessing purity

A mysterious white powder, a blood smear, and a half eaten biscuit – completely unrelated items to most, but at a crime scene they are important evidence for forensic scientists to analyse. Forensic scientists use chemical methods to analyse traces of crime scene material – gunshot residues, drugs, hair or blood. Chemical analysis is also used to separate and test the materials in our food and water, and to identify and test the purity of different substances.

Pure substances

When we see the stamp 'pure orange juice', like that shown in Figure 6.1, we consider the orange juice to be a natural substance with nothing added to it during manufacture; just the juice without colourings or sweeteners. A chemist, however, would not consider the orange juice to be pure as it is a mixture of different substances.

In chemistry **a** pure **substance is a single element or compound not mixed with any other substance**.

Orange juice is not a single element or compound, as it contains water, citric acid, vitamin C and other ingredients. It is a mixture and not a pure substance.

Figure 6.1 A scientist would not consider orange juice to be pure.

Table 6.1 Some pure substances and some mixtures

Pure substances	Mixtures
Diamond only contains the element carbon.	Air is a mixture of oxygen, carbon dioxide, nitrogen and other trace gases.
Water only contains the compound water.	Mineral water is a mixture of water and dissolved salts.
Table salt only contains the compound sodium chloride.	Milk is a mixture of water, lactose, fat and minerals.

Figure 6.2 Mineral water is not pure as it is not a single compound. It is mostly water but there are other substances mixed with it. These are the ingredients that you see listed on the label.

Melting and boiling point of pure substances and mixtures

Tip

The melting point of a substance is the same temperature as the freezing point.

Tips are given throughout to aid understanding.

Melting point is the temperature at which a solid changes into a liquid.

Boiling point is the temperature at which a liquid changes into a gas.

Pure substances have specific melting points and boiling points. For example, the compound water has a boiling point of 100 °C and a melting point of 0 °C; the element sulfur has a melting point of 115 °C and a boiling point of 444 °C. Substances can be identified by their melting points and boiling points.

A substance must absorb heat energy so that it can melt or boil. The temperature of the substance does not change during melting, boiling or freezing, even though energy is still being transferred. You can see this in the graph in Figure 6.3.

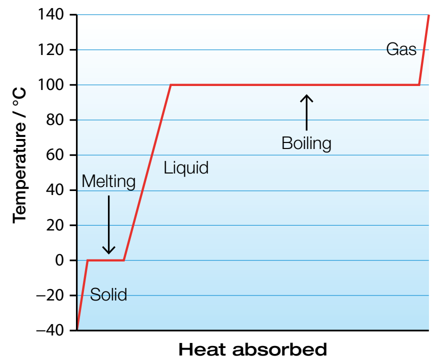


Figure 6.3 Graph of temperature against heat absorbed for water. Note that the temperature is constant during melting and boiling

Impure substances do not have sharp melting or boiling points, but melt or boil over a range of temperatures. The presence of an impurity usually lowers the melting point of a solid or raises the boiling point of a liquid. The greater the amount of an impurity, the bigger the difference from the true melting point and boiling point.

Table 6.2 The differences in melting and boiling points between pure and impure substances

Pure substances	Mixtures (impure substances)
Have sharp, specific melting and boiling points	Melt or boil over a range of temperatures
	An impure substance has a lower melting point and a higher boiling point

Tip

Remember that a formulation is a mixture; no chemical reaction occurs when it is made.

Formulations

A formulation is a mixture that has been designed as a useful product. It is formed by mixing together several different substances in carefully measured quantities to ensure the product has the required properties.

Many foods are formulations and this means they have consistent properties. Would you buy tomato ketchup if it tasted different each time? An 18-carat-gold alloy used in jewellery is a formulation; it is a mixture of carefully measured quantities of gold, silver, copper and zinc to ensure that it has the correct hardness. Three types of formulation are shown in Table 6.3.

Table 6.3 Examples of formulations

Formulation	Comment	Named example
Alloys	An **alloy** is a mixture of two or more elements, at least one of which is a metal. The resulting mixture has metallic properties. Each alloy is a formulation made by mixing measured quantities to ensure the alloy has the specific properties required for its use.	18-carat gold (gold, silver, copper, zinc)
Medicine	Tablets are formulations of the active drug along with ingredients such as corn starch to bind the tablet together and a lubricant to make it easy to swallow. Liquid medicines are also formulations.	Calpol® is a formulation of paracetamol and liquid flavourings.
Fertiliser	Fertilisers contain mixtures of nitrogen, phosphorus and potassium compounds. Different formulations are suitable for different plants.	NPK fertilisers

Consolidate learning with regular Test Yourself questions.

Test yourself

1 **a)** Why do some people state that a bottle of mineral water is pure water?
b) Explain why a chemist would not say that mineral water is pure.
c) Name all the pure substances from the following list: apple juice, copper, sodium chloride, rock salt, air, magnesium, milk.

2 **a)** A sample of water was found to freeze between –1 °C and –3 °C. Was this water pure? Explain your answer.
b) Pure water boils at 100 °C. What happens to the boiling point if some salt is dissolved in the water?

3 The melting point of paracetamol is 169 °C. Four students, A–D, made paracetamol in the laboratory and recorded the melting point. Which student(s) made pure paracetamol? Explain your answer.

Student	A	B	C	D
Melting point/°C	164–166	160–162	169	169–170

4 The melting point of aspirin is 136 °C. A student prepared some aspirin and found the melting point to be 132–134 °C. What can the student deduce about the prepared aspirin?

5 A fertiliser contained 23% ammonium nitrate, 23% magnesium sulfate and 54% ammonium phosphate. Why is this fertiliser a formulation?

Show You Can questions are designed for class discussion or to enhance individual learning.

Show you can

The table shows the melting points and boiling points of some metallic elements and alloys named by the letters A–C.

	Melting point/°C	Boiling point/°C
A	–34	356
B	420	913
C	1425–1540	2530–2545

a) In what state is substance A at room temperature and pressure (20 °C)?
b) In what state is substance B at room temperature and pressure (20 °C)?
c) In what state is substance C at room temperature and pressure (20 °C)?
d) What is an alloy?
e) Classify the substances in the table as elements or alloys and explain your answer.

Separating mixtures

A mixture is defined as two or more substances mixed together. The substances in a mixture are quite easy to separate because the substances are not chemically joined to each other. It is useful to understand the following key terms when thinking about separating mixtures.

- A **solute** is the substance that dissolves in a solvent.
- A **solvent** is the liquid in which a solute dissolves.
- A **solution** is a solute dissolved in a solvent.
- A **soluble substance** is one which will dissolve in a solvent.
- An **insoluble substance** is one which does not dissolve in a solvent.

Different methods are used to separate different types of mixtures.

1 Filtration

Filtration is used to separate an insoluble solid from a liquid. For example, it could be used to separate solid insoluble sand from water.

The mixture is poured through a filter funnel containing a piece of filter paper. The insoluble solid is caught in the filter paper but the liquid passes through. The liquid which passes through the filter paper during filtration is called the **filtrate**. The solid that remains on the filter paper after filtration is called the **residue**.

Tip

When drawing diagrams, always make sure they are labelled.

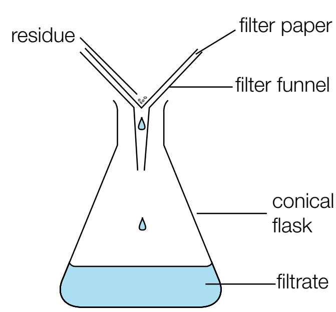


Figure 6.4 Filtration

2 Evaporation

To separate a dissolved solid from a solvent, the mixture is placed in an evaporating basin and heated gently until all the solvent has evaporated and the solid is left in the basin. This method is often used to separate salt from a salt solution.

Evaporation **is the change of state from liquid to gas when heated.**

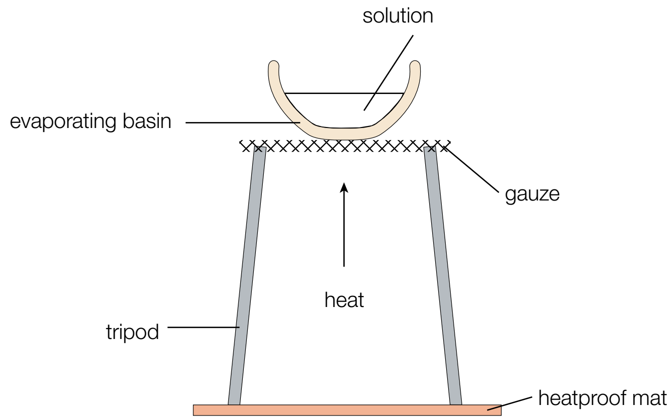


Figure 6.5 Evaporation

3 Crystallisation

Another method of separating a dissolved solid from a solvent involves heating the solution to boil off some of the solvent. This creates a saturated solution. **A** saturated solution **is one in which no more solid can dissolve at that temperature**. The saturated solution is then cooled. The dissolved solid (solute) becomes less soluble and so cannot remain dissolved. Consequently, it crystallises out of solution. The crystals may then be separated from the saturated solution by filtration. Many salts, for example copper(II) sulfate, are separated from their solution by crystallisation.

Figure 6.6 Crystals of copper(II) sulfate produced by crystallisation

Tip

Check out Chapter 7 for more information on saturated solutions and solubility.

4 Simple distillation

To separate a solvent from a solution, simple distillation can be used. For example, pure water can be obtained from sea water or from copper(II) sulfate solution using distillation. Distillation is evaporation followed by condensation.

Condensation **is the change of state from gas to liquid when cooled.**

The mixture to be separated is placed in the distillation apparatus, as shown in Figure 6.7, and heated until the solvent boils. Anti-bumping granules are added to the mixture in the flask to promote smooth boiling.

The vaporised solvent passes into the condenser where it cools and condenses and then runs into the collection flask. **The** distillate **is the liquid that is cooled from the vapour and collected during distillation.**

Clear and colourful diagrams help develop knowledge.

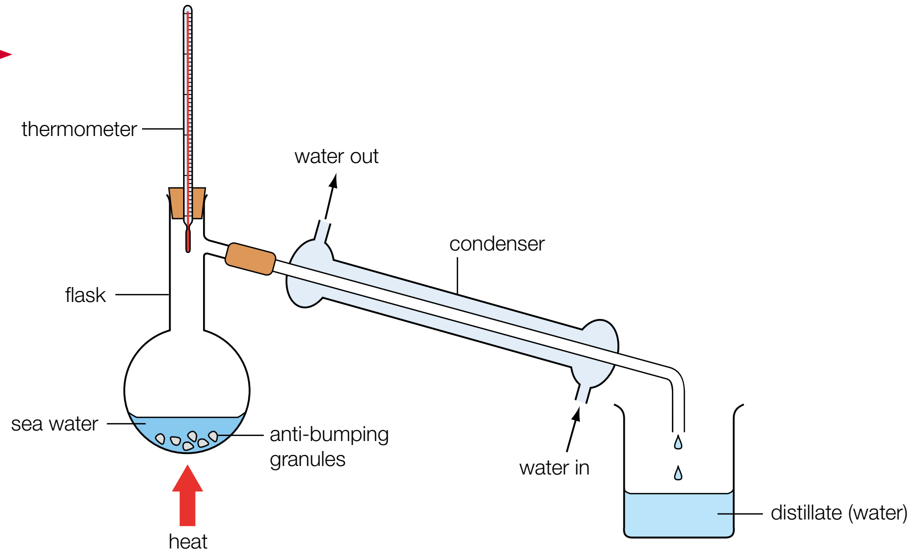


Figure 6.7 Simple distillation of sea water to obtain pure water

Tip

Remember when labelling that water enters a condenser at the bottom and leaves at the top; this makes sure the outer jacket is completely filled and an effective cooling system is created.

Tip

The bulb of the thermometer is opposite the exit to the condenser so that the temperature of the exiting vapour, which condenses to form the product, can be recorded.

5 Fractional distillation

Fractional distillation is used to separate miscible liquids which have different boiling points.

Miscible liquids are liquids that mix together, such as alcohol and water.

Immiscible liquids are liquids which do not mix together but form two distinct layers, for example oil and water.

Figure 6.8 Gin and tonic water are miscible liquids

Figure 6.9 Oil and water are immiscible

The mixture of miscible liquids is heated and the liquids boil one by one as the temperature rises. Each vapour rises up, is condensed and then collected. The distillate collected at each different temperature is called a **fraction** and a different receiver is used for each fraction.

When separating a mixture of ethanol and water, the ethanol boils at 79 °C and is condensed and collected first. The temperature then rises to 100 °C, at which point water boils and is collected.

As you can see in Figure 6.10, the apparatus used is similar to that for simple distillation, but a long column called a **fractionating column** is used; this provides a better separation of liquids. Any evaporated liquids below their boiling point condense on the glass beads in the fractionating column and run back into the flask. This means the fractions are pure.

Tip

If the liquids in the mixture are flammable, it is safer to heat the flask with an electric heating mantle or a water bath (if the temperature needed is less than 100 °C), rather than a Bunsen burner.

Tip

You will study fractional distillation of crude oil in industry in Unit 2.

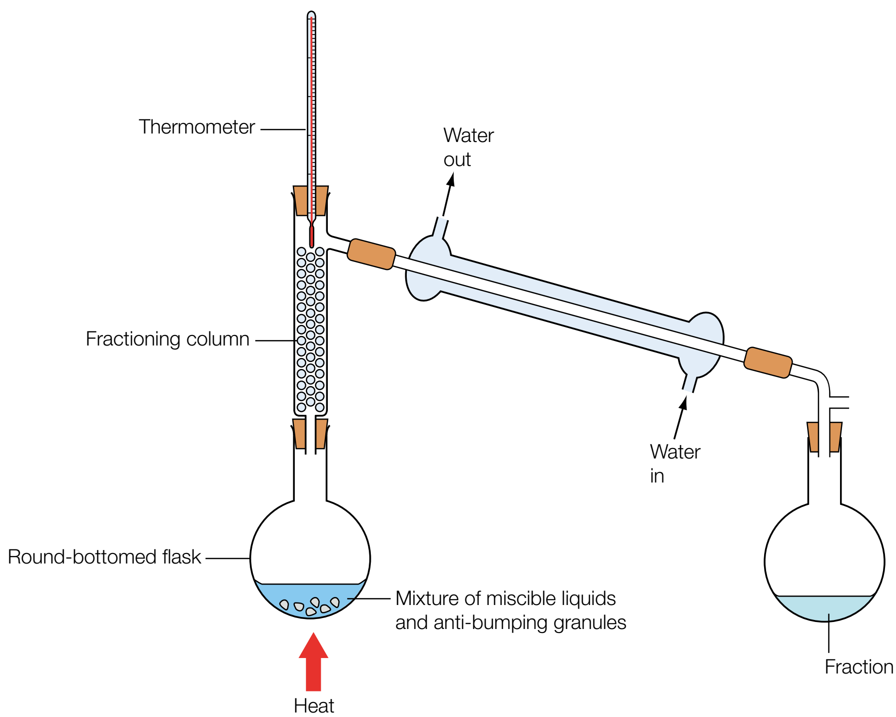


Figure 6.10 Fractional distillation apparatus

Test yourself

6 A mixture of insoluble copper(II) carbonate and water was separated by filtration.
   a) Name the filtrate.
   b) Name the residue.
7 Pure water can be obtained by distillation of copper(II) sulfate solution.
   a) Name the solute in copper(II) sulfate solution.
   b) Name the change of state which occurs in the distillation flask.
   c) What is the distillate?
   d) What is added to the distillation flask to ensure even boiling?

Show you can ?

Three common methods of separation are filtration, distillation and fractional distillation.

For each of these separation methods pick two words or phrases from the list below and insert them into a copy of the table with an explanation of their meaning.

condenser, distillate, fractionating column, filtrate, miscible liquids, residue

Complete the table by writing the type of mixture separated by each method.

	Filtration	Distillation	Fractional distillation
Type of mixture separated			
Important word and definition			
Important word and definition			

Practical activity

Rock salt

Common salt has the chemical name sodium chloride. It is found naturally in large amounts in seawater or in underground deposits. Sodium chloride can be extracted from underground deposits by the process of solution mining, which is shown in Figure 6.11.

1 a) On what physical property of sodium chloride does this process depend?
   b) Suggest one reason why solution mining uses a lot of energy.
   c) Suggest one negative effect that solution mining has on the environment.
   d) How is sodium chloride obtained from the concentrated salt solution?

Rock salt is a mixture of salt, sand and clay. Pure salt can be separated from rock salt using the laboratory method detailed below.

Method

a) Place eight spatulas of rock salt into a mortar and grind using a pestle.
b) Place the ground rock salt into a beaker and quarter fill with water.
c) Place the beaker on a gauze and tripod and heat gently while stirring with a glass rod. Stop heating when the salt has dissolved. The sand and clay will be left.
d) Allow the mixture to cool and then filter into an evaporating basin.
e) Heat until half the volume of liquid is left.
f) Place the evaporating basin on the windowsill to evaporate off the rest of the water slowly. Pure salt crystals should be left.

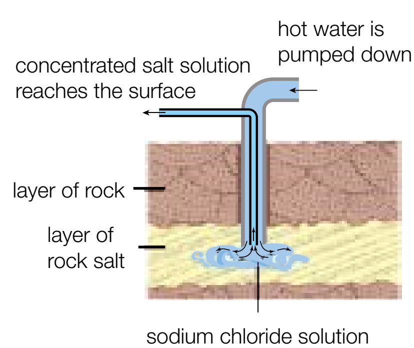


Figure 6.11 Salt extraction by solution mining

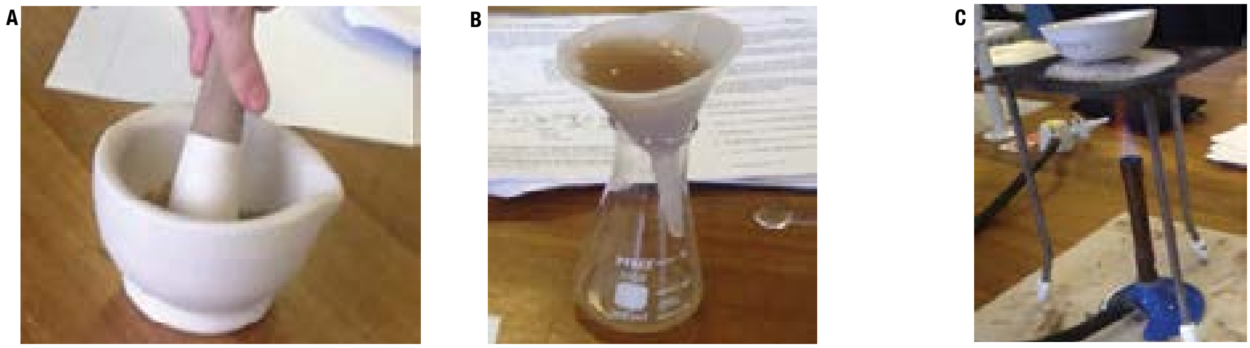


Figure 6.12

2 Which step of the method detailed on page 11 (a–f) is best represented by each photograph in Figure 6.12?

3 a) Why is rock salt considered to be a mixture?
  b) What was the purpose of grinding the rock salt?
  c) Why was the mixture heated and stirred?
  d) State what the filtrate contains.
  e) State what the residue contains.
  f) Explain why the salt obtained may still be contaminated with sand and suggest how you would improve your experiment to obtain a purer sample of salt.

6 Paper chromatography

Chromatography can be used to separate mixtures of soluble substances in a solution. Mixtures that are suitable for separation by chromatography include inks, dyes and colouring agents in food.

Figure 6.13 Paper chromatography of four different inks

Method

1. Using a pencil, draw a base line on the chromatography paper around 1-2 cm from the bottom. Use pencil as it will not dissolve in the solvent.
2. Using a capillary tube, place a spot of the substance to be analysed on the base line. When this is dry, add another spot on top to make it concentrated.
3. Hang the paper in a beaker with the solvent at the bottom. The pencil line and spots must be above the level of the solvent so that the spots do not dissolve into the solvent in the beaker.
4. The solvent travels up the paper.
5. When the solvent is near the top, take the paper out of the solvent and mark the level that the solvent reached. This is known as the solvent front.
6. Leave the paper to dry. The mixture should have separated into different components, which are seen as spots on the paper.

Tip

If the spots are colourless and not visible, they can be sprayed with a chemical developing agent.

Solvent front is the furthest distance travelled by the solvent.

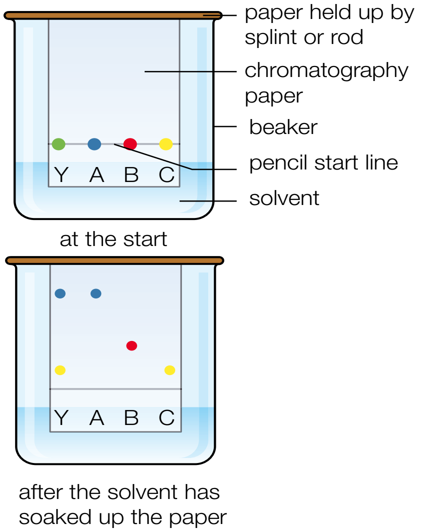


Figure 6.14 Paper chromatography

Tip

In an examination you will need to use a ruler to measure the distance moved by the substance and the solvent. If the spot is large, measure to the middle of the spot.

A pure substance usually produces a single spot in chromatography, whatever solvent is used. In Figure 6.14, A, B, and C are pure substances. Mixtures will usually produce more than one spot; one for each substance in the mixture. Y is a mixture of two substances as it produces two spots. By comparing Y with A, B and C, it should be clear that Y is a mixture of substances A and C.

How chromatography works

In paper chromatography, the **stationary phase** is the sample mixture on the paper and the **mobile phase** is the solvent which runs through the stationary phase.

Each substance in the mixture moves at a different rate up the paper. This depends on its relative attraction to the paper and the solvent. Substances that have a stronger attraction to the solvent, and are more soluble in it, move quickly and travel a long way up the paper. Substances that have a stronger attraction to the paper move slowly and only travel a short distance up the paper (for example substance C in Figure 6.14).

R_f values

The separated components of the mixtures can be identified by calculating the R_f value using the equation:

$$R_f = \frac{\text{distance moved by substance}}{\text{distance moved by solvent}}$$

The R_f value of a particular substance in the same solvent is always the same. An example of how to calculate R_f values for two substances, P and Q, is shown in Figure 6.15.

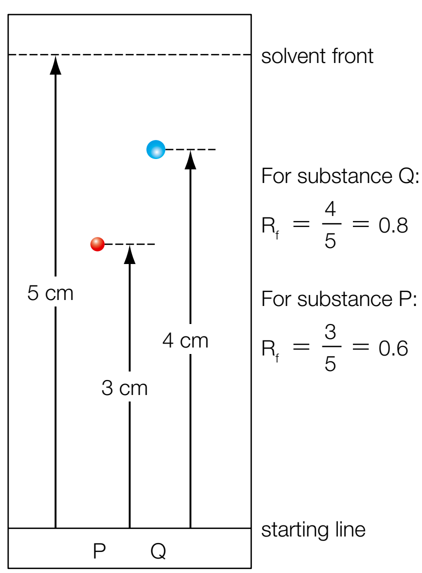


Figure 6.15 Finding R_f values

Test yourself

8 The chromatogram in Figure 6.16 shows the results when a mixture of dyes, S, was compared to substances 1–6.
  a) Explain which of substances 1 to 6 are pure substances.
  b) Identify the substances in S.
  c) Calculate the R_f values for each spot in colouring S. Give your answer to two significant figures.
  d) Which colour spot moved slowest during the experiment?
  e) Explain what is meant by the terms solvent and solvent front.
  f) Name the stationary phase.
  g) When setting up this experiment how would you draw the base line?

9 The same substance was analysed by chromatography using two different solvents, as shown in Figure 6.17.
  a) Calculate the R_f value of the substance in each solvent.
  b) Explain why the substance moved further in solvent 2 than solvent 1.

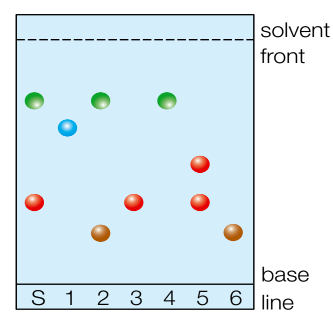


Figure 6.16

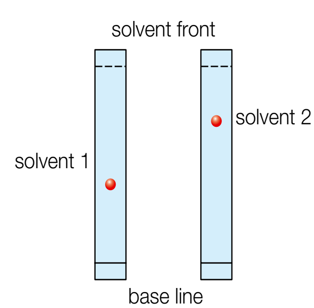


Figure 6.17

Show you can

Name the best method of separation in each case:

a) to separate the dyes in red food colouring
b) to separate salt and water from salt solution
c) to produce crystals of copper(II) sulfate from copper(II) sulfate solution
d) to separate a miscible mixture of wine and water
e) to separate water from copper(II) sulfate solution
f) to separate insoluble copper(II) oxide and water
g) to separate a miscible mixture of ethanoic acid, water and ethanol.

F

Choosing methods of separation

To decide on the most appropriate method of separating a mixture, it can be useful to first decide the type of mixture. The summary table below can then help you decide which method is best.

Table 6.4 Summary of separation methods

Type of mixture	Insoluble solid and liquid	Soluble solid dissolved in a solvent	Two miscible liquids	Soluble solids dissolved in a solvent (often coloured)
Example	Sand and water	Sodium chloride solution	Ethanol and water	Dyes Food colourings Ink
Method of separation	Filtration	Evaporation (to obtain solid) Crystallisation (to obtain solid) Simple distillation (to obtain solvent) chromatography		Paper chromatography

Water

Types of water

Water is essential for life. To ensure your body functions well, you need to drink plenty of water every day. The recommended amount is at least a litre, though this value depends on physical activity, age and environmental conditions. Water that is safe to drink is called potable water. Water covers about 70% of the Earth's surface. There are many different types of water and these are given in Table 6.5.

Table 6.5 Types of water

Type	Description	Dissolved salts	Are microbes present?
Pure water	Water which contains water molecules only	None	No
Fresh water	Water which is naturally occurring in lakes, ice caps, rivers and underground rocks	Low levels	Very few
Groundwater	Fresh water found underground in soil spaces and in porous rocks	Low levels	Yes
Sea water	Water in the sea and oceans	Low levels	Yes
Waste water	Used water from homes, industry and agriculture	Low levels	Yes

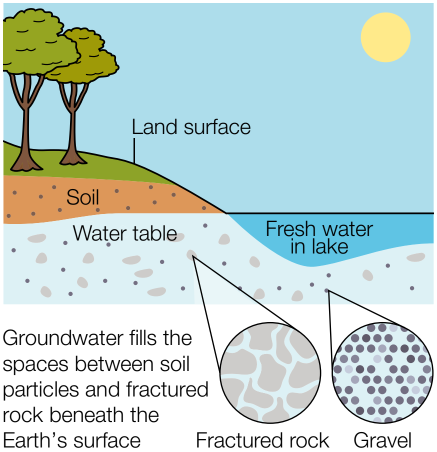


Figure 6.18

Making water potable

Potable water is obtained from a number of different sources.

1 Potable water from fresh water

In the UK, potable water is made from fresh water which comes from rain and collects in lakes, rivers, reservoirs and also in porous rocks underground (groundwater, shown in Figure 6.18). This water may contain dissolved salts and minerals, microbes, pollutants and insoluble materials like sand and stones. It must be treated before it is fit to drink. The stages in making potable water from fresh water are shown in Table 6.6.

Table 6.6 The stages in water treatment

Process	What it does	Description
Filtration	Removes insoluble solids e.g. stones, soil, leaves	The water is sprayed onto specially prepared layers of sand and gravel called filter beds. Different-sized insoluble solids are removed as the water trickles through the filter beds; the sand and gravel act like filter paper. The filter beds are cleaned every so often by pumping clean water backwards through the filter.
Sedimentation	Clumps tiny particles together into large particles which settle out	In a sedimentation tank, aluminium sulfate is added to clump tiny particles together to make larger particles, which settle out more easily. The water is then passed through a fine filter, such as carbon granules, to remove very small particles.
Chlorination	Kills microbes	Chlorine gas is bubbled into the water to kill microbes and sterilise the water.

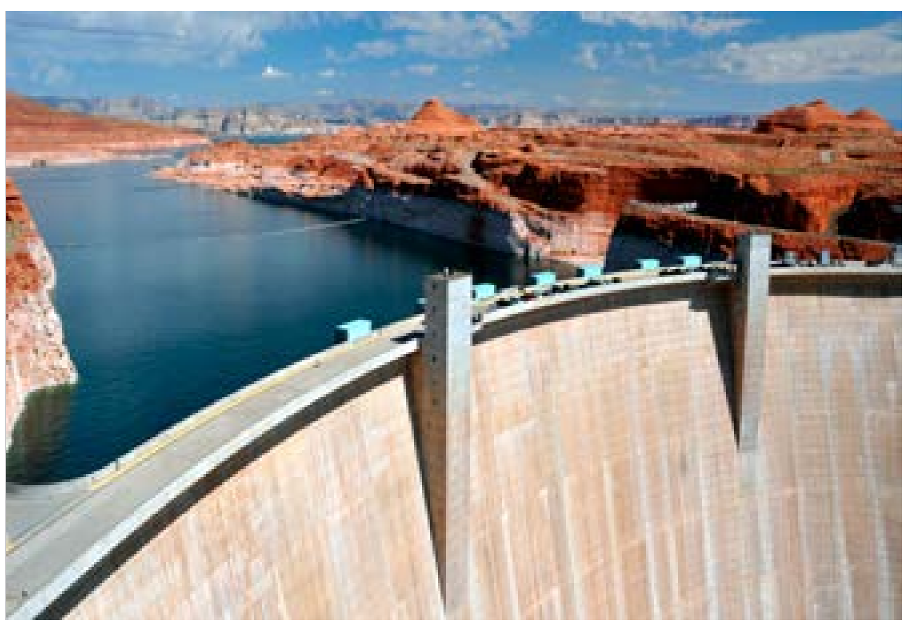


Figure 6.19 Lake Mead is an artificial reservoir formed by damming the Colorado River at the Hoover Dam. It is the largest reservoir in the USA and supplies fresh water to Las Vegas and beyond.

2 Potable water from waste water

Waste water from domestic, industrial and agricultural sources contains organic matter, microbes and harmful chemicals. Waste water may also include sewage and must be treated before it can be returned to the environment. The stages in the treatment are shown in Figure 6.20.

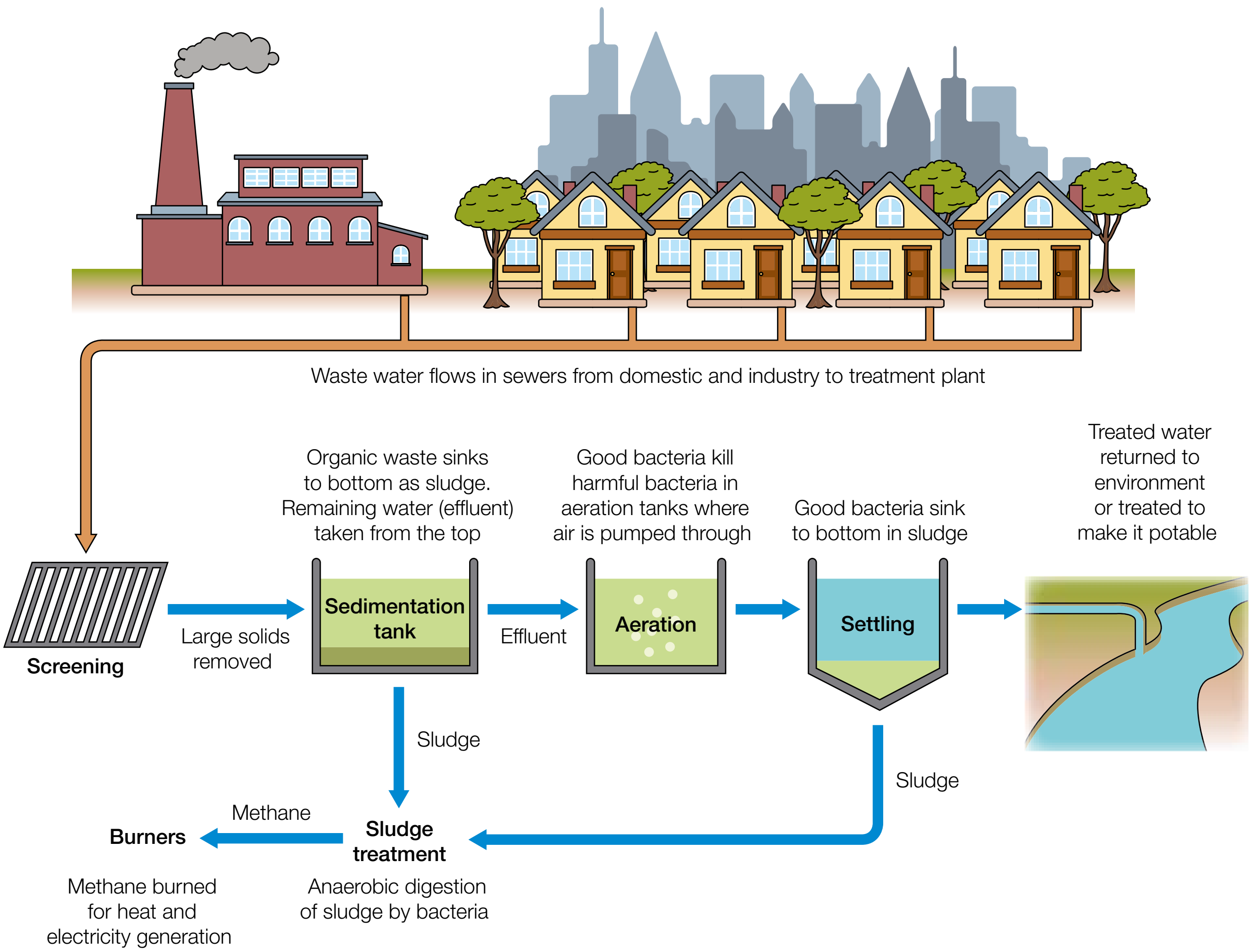


Figure 6.20 Waste water treatment

The different steps include

- **screening**: this removes large solids
- **sedimentation**: the heavy organic human waste sinks to the bottom as sludge; this separates it from the rest of the water, which is called the effluent
- **aerobic treatment of effluent**: air is passed through the effluent in aeration tanks which leads to good bacteria killing harmful bacteria.

The treated water is then returned to the environment, or it goes through filter beds and is chlorinated to make it potable.

The sludge can be digested by bacteria in the absence of air (anaerobic) to produce methane which can be burned for heat and electricity generation.

Figure 6.21 A desalination plant at Lake Mead, Nevada, USA. Due to the increasing population of Las Vegas and drought conditions, desalination has been introduced to meet the water needs of the city.

3 Potable water from seawater

Cruise ships use more than 260 000 gallons of water every day. Rather than carrying all this water from the various ports, newer ships transform salty seawater into potable water by a process known as desalination. **Desalination is the process of removing dissolved substances from sea water**. Desalination is often carried out by distillation. The sea water is heated so that it boils; the water evaporates and turns to steam, leaving behind the dissolved substances. The pure water is cooled in a condenser.

Desalination by distillation requires large amounts of energy to boil the water and as a result it is a very expensive process. Many hot countries produce potable water by distillation as they can use solar power to reduce costs. Saudi Arabia produces 70% of its drinking water by desalination as it has an abundance of cheap oil which can produce the energy required economically.

Test for water

Figure 6.22 White anhydrous copper(II) sulfate turns blue in the presence of water.

Anhydrous copper(II) sulfate ($CuSO_4$) does not contain any water of crystallisation and is powdery. When water is added, blue hydrated copper(II) sulfate forms, which is crystalline. This is an exothermic reaction.

The equation shows that hydrated copper(II) sulfate contains five moles of water of crystallisation.

anhydrous copper(II) sulfate + water → hydrated copper(II) sulfate

$$CuSO_4 + 5H_2O \rightarrow CuSO_4.5H_2O$$

This reaction can be used as a chemical test for the presence of water. A few drops of the liquid to be tested are added to white anhydrous copper(II) sulfate. If water is present, the white powder will turn blue.

Tip

Remember that hydrated means contains water of crystallisation and anhydrous means contains no water of crystallisation.

Test yourself

10 In the UK, potable water is produced from fresh water.
- **a)** What is potable water?
- **b)** What is fresh water?
- **c)** Key steps in the production of potable water are filtration, sedimentation and chlorination.
  - **i)** Why and how is filtration carried out?
  - **ii)** Name a chemical used for sedimentation and state why sedimentation is carried out.
  - **iii)** Why is chlorination carried out?

11a) What is desalination?
- **b)** More potable water is produced by desalination in Saudi Arabia than in any other country. Why is water produced this way in Saudi Arabia?
- **c)** Describe how potable water is produced from sea water by distillation.

12a) Give three sources of waste water.
- **b)** Describe what happens at each of the following stages of sewage treatment.
  - **i)** screening
  - **ii)** sedimentation
  - **iii)** aerobic treatment of effluent
  - **iv)** anaerobic treatment of sludge.

13 What is the chemical test for water?

Show you can ?

	Waste water	Ground water	Salt water
Method of producing potable water			Desalination by distillation

In 2010, the Thames Water desalination plant, the first and only desalination plant in the UK, opened. It is able to produce 1.5% of the UK's water requirement. Israel has many desalination plants and 40% of its water requirement is supplied in this way. Singapore, once heavily dependent on Malaysia for its water supply, now recycles sewage and waste water and uses this as a primary supply of potable water.

a) Copy and complete the table to give details of the methods of converting the different types of water into potable water. One column has been completed for you.

b) State and explain which method of obtaining potable water is primarily used in the UK.

c) Suggest reasons why there is a difference in the percentage use of desalination in the UK and in Israel.

d) Suggest why using waste water has been important for Singapore.

Figure 6.23 Flame colours of different ions can be observed when fireworks burn. We can work out the metal ions present in the fireworks shown.

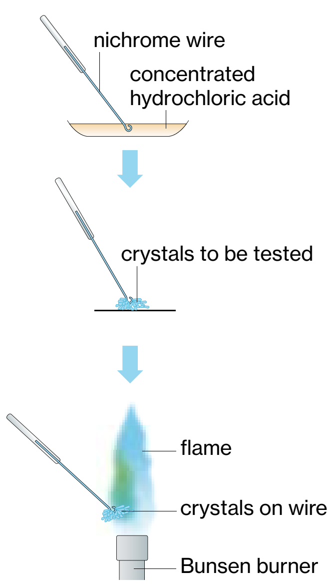


Figure 6.24 Procedure for a flame test

Tests for ions

Salts are ionic compounds, which contain positive ions and negative ions. Positive ions are called cations and negative ions are called anions. Table 6.7 shows the cations and anions found in some different salts.

Table 6.7 The cations and anions in some common salts.

Salt	Cation	Anion
Sodium chloride	Sodium (Na^+)	Chloride (Cl^-)
Magnesium sulfate	Magnesium (Mg^{2+})	Sulfate (SO_4^{2-})
Copper(II)carbonate	Copper(II) (Cu^{2+})	Carbonate (CO_3^{2-})

It is possible to identify the ions present in a salt by carrying out cation tests and anion tests.

Cation tests

Cations can be identified in two different ways: flame tests or the addition of sodium hydroxide solution or ammonia solution.

1 Flame tests

Some positive metal ions produce an intense colour in a blue Bunsen burner flame. A flame test can be used to identify them.

Method

- Make a loop on the end of a piece of nichrome wire.
- Dip the loop into concentrated hydrochloric acid and then into the salt to be tested.
- Place the loop into a blue Bunsen burner flame and record the first colour observed.

The flame test colours which you need to learn are shown in Table 6.8.

Figure 6.25 The photograph shows a flame test being carried out on a potassium salt.

Table 6.8 Flame test colours

Metal ion present	Flame colour
Lithium (Li^+)	Crimson
Sodium (Na^+)	Yellow/orange
Potassium (K^+)	Lilac
Calcium (Ca^{2+})	Brick-red
Copper (Cu^{2+})	Blue-green/green-blue

Test yourself

14 Describe how you would determine experimentally whether a solid powder A contained potassium ions.

15 Write the formula of the metal ion which causes
a) a yellow flame colour
b) a blue-green flame colour.

16 What is a cation?

17 What is observed when a flame test is carried out on
a) calcium chloride
b) lithium chloride?

18 Copy and complete the table.

Ion name	Ion formula	Cation or anion?
	Na^+	
	HCO_3^-	
Iron(II)		
Aluminium		
Sulfate		
	Cl^-	
	CO_3^{2-}	
Iodide		
Copper(II)		

Tip

Flame colour is produced by the metal ion in the compound. Don't confuse flame colours with the colour of the flame when a metal burns in air. Sometimes the colour is the same, but in others it is not. For example, magnesium compounds do not have a flame test colour but magnesium burns with a white flame in air.

2 Test using sodium hydroxide solution or ammonia solution

F

When a metal ion solution is mixed with sodium hydroxide solution or ammonia solution, **precipitates** are produced. These can be used to identify the metal cation.

A precipitate is a solid formed when two solutions are mixed.

For example, when solutions of iron(II) sulfate and sodium hydroxide are mixed, a green precipitate is formed. In some cases the precipitate will dissolve if excess sodium hydroxide or ammonia solution is added.

Method

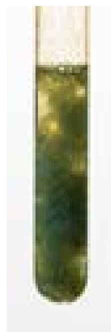
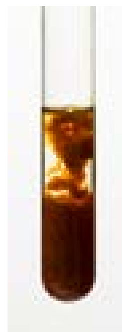
- If the salt to be tested is a solid, make a solution by dissolving a spatula of the solid in water.
- Place about 5 cm^3 of the metal ion solution into a test tube and add a few drops of sodium hydroxide solution.
- Record the colour of the precipitate.
- Continue to add sodium hydroxide solution until it is in excess and record the effect.

Tip

It is acceptable to shorten the word precipitate to ppt.

Ammonia solution can be used instead of sodium hydroxide solution as the results are the same. For aluminium and copper(II) salts, however, the results on adding excess sodium hydroxide and ammonia solution are different. These are shown in bold in Table 6.9.

Table 6.9 Testing for cations with ammonia solution and sodium hydroxide solution

Metal ion	Copper(II) Cu^{2+}	Iron(II) Fe^{2+}	Iron(III) Fe^{3+}	Magnesium Mg^{2+}	Aluminium Al^{3+}	Zinc Zn^{2+}
Result on adding dilute sodium hydroxide or ammonia solution	Blue ppt of copper(II) hydroxide 	Green ppt of iron(II) hydroxide 	Brown ppt of iron(III) hydroxide 	White ppt of magnesium hydroxide 	White ppt of aluminium hydroxide 	White ppt of zinc hydroxide 
Ionic equation	$Cu^{2+}(aq) + 2OH^-(aq) \rightarrow Cu(OH)_2(s)$	$Fe^{2+}(aq) + 2OH^-(aq) \rightarrow Fe(OH)_2(s)$	$Fe^{3+}(aq) + 3OH^-(aq) \rightarrow Fe(OH)_3(s)$	$Mg^{2+}(aq) + 2OH^-(aq) \rightarrow Mg(OH)_2(s)$	$Al^{3+}(aq) + 3OH^-(aq) \rightarrow Al(OH)_3(s)$	$Zn^{2+}(aq) + 2OH^-(aq) \rightarrow Zn(OH)_2(s)$
Effect of excess sodium hydroxide solution	**Blue ppt remains**	Green ppt remains	Red-brown ppt remains	White ppt remains	**White ppt dissolves and a colourless solution is produced**	White ppt dissolves and a colourless solution is produced
Effect of excess ammonia solution	**Blue ppt dissolves and a deep blue solution is produced**	Green ppt remains	Brown ppt remains	White ppt remains	**White ppt remains**	White ppt redissolves and a colourless solution is produced

Note that using sodium hydroxide solution alone to identify zinc ions positively is not a valid test because the same result is obtained for aluminium ions: a white ppt which dissolves on excess to give a colourless solution. The test is not conclusive for these ions and ammonia solution needs to be used to identify the zinc ions positively.

To test for aluminium ions, first add sodium hydroxide solution and then retest with ammonia solution, as aluminium ions give the same result as magnesium ions with ammonia solution, and the same result as zinc ions with excess sodium hydroxide solution.

What happens in the test?

When the two solutions are mixed, a **precipitation reaction** occurs; the positive ions from one solution react with the negative ions of the other and an insoluble solid precipitate forms.

The sodium hydroxide solution and the ammonia solution contain hydroxide ions. The negative hydroxide ions react with the positive metal ions to form a metal hydroxide. Metal hydroxides (except for those in Group 1) are insoluble, so a precipitate forms.

When solutions of copper(II) sulfate and sodium hydroxide are mixed, a blue precipitate of copper(II) hydroxide is formed, as the positive copper(II) ions from the copper(II) sulfate react with the negative hydroxide ions of the sodium hydroxide solution.

Word equation:

copper(II) sulfate + sodium hydroxide → copper(II) hydroxide + sodium sulfate

Symbol equation: $CuSO_4(aq) + 2NaOH(aq) \rightarrow Cu(OH)_2(s) + Na_2SO_4(aq)$

Ionic equation: $Cu^{2+}(aq) + 2OH^-(aq) \rightarrow Cu(OH)_2(s)$

The sulfate ions and the sodium ions are spectator ions and take no part in the reaction; they are left behind in the solution as sodium sulfate. The ionic equation gives a full picture of the reaction which is occurring.

When solutions of iron(III) nitrate and sodium hydroxide are mixed, a brown precipitate of iron(III) hydroxide is formed as the positive iron(III) ions from the iron(III) nitrate react with the negative hydroxide ions of the sodium hydroxide.

Word equation:

iron(III) nitrate + sodium hydroxide → iron(III) hydroxide + sodium nitrate

Symbol equation: $Fe(NO_3)_3(aq) + 3NaOH(aq) \rightarrow Fe(OH)_3(s) + 3NaNO_3(aq)$

Ionic equation: $Fe^{3+}(aq) + 3OH^-(aq) \rightarrow Fe(OH)_3(s)$

The nitrate ions and the sodium ions are spectator ions in this case.

Tip

If you are unsure about writing balanced symbol and ionic equations, check out page 20.

Tip

You do not need to write full symbol equations for the reaction of ammonia solution with metal ion solution. You do, however, need to know the ionic equations, as shown in Table 6.9.

Test yourself

19 Iron ions are found in garden moss killer. Describe how you would determine experimentally whether a moss killer contained iron(II) ions or iron(III) ions.

20 Describe how you would carry out a valid test to determine whether a solid contained aluminium ions.

21 a) What is observed if excess ammonia solution is added to a solution of
  - **i)** copper(II) sulfate
  - **ii)** magnesium nitrate?

b) Write ionic equations, with state symbols, for the reactions which occur in (a).

c) Are the observations in part (a) different if sodium hydroxide solution is used instead of ammonia solution?

22 Five drops of sodium hydroxide solution were added to 2 cm^3 of each of the metal ion solutions below. Write down the observations which occur in each case and write a balanced symbol equation. Write ionic equations for the reactions which occur.

a) copper(II) nitrate
b) magnesium sulfate
c) iron(II) chloride
d) aluminium nitrate

23 If some sodium hydroxide solution was added to a solution of potassium chloride, would a precipitate form? Explain your answer.

Anion tests

1 Test for halide ions

Halide ions are the anions (negative ions) formed from atoms of Group 7 elements (the halogens). They include chloride (Cl^-), bromide (Br^-) and iodide (I^-).

> **Tip**
> A common error is to write 'if a white ppt forms, the solid contains chlorine'. This is not correct; the white ppt indicates the presence of chlor*ide ions*.

Method

- If the sample is a solid, dissolve it in water; if insoluble in water, use nitric acid.
- Add a few drops of silver nitrate solution.
- If a white ppt forms, the solid contains chloride ions. If a cream ppt forms, the solid contains bromide ions. If a yellow ppt forms, the solid contains iodide ions.

Table 6.10 shows the results for this test.

Table 6.10 Tests for halide ions

Halide ion	Chloride, Cl^-	Bromide, Br^-	Iodide, I^-
Result on adding silver nitrate solution	White ppt of silver chloride 	Cream ppt of silver bromide 	Yellow ppt of silver iodide 
Ionic equation	$Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$	$Ag^+(aq) + Br^-(aq) \rightarrow AgBr(s)$	$Ag^+(aq) + I^-(aq) \rightarrow AgI(s)$

> **Tip**
> Remember that aqueous means a solution in water, so silver nitrate solution is often referred to as aqueous silver nitrate.

What happens in the test?

The negative halide ions react with the positive silver ions from the silver nitrate solution to form a silver halide: silver chloride, bromide or iodide. Silver halides are insoluble so a precipitate forms. This is a precipitation reaction.

For example, when solutions of sodium bromide and silver nitrate are mixed, a cream precipitate of silver bromide is formed as the positive silver ions from the silver nitrate react with the negative bromide ions of the sodium bromide.

Word equation:
sodium bromide + silver nitrate → silver bromide + sodium nitrate

Symbol equation: $NaBr(aq) + AgNO_3(aq) \rightarrow AgBr(s) + NaNO_3(aq)$

Ionic equation: $Ag^+(aq) + Br^-(aq) \rightarrow AgBr(s)$

> **Tip**
> A reagent is a chemical used to create a reaction in chemical analysis. Sodium hydroxide solution, silver nitrate solution and barium chloride solution are all reagents.

2 Test for sulfate ions

Method

- If the sample is a solid, dissolve it in water; if insoluble in water, use nitric acid.
- Add a few drops of barium chloride solution.
- If a white ppt forms, the solid contains sulfate ions.

> **Tip**
> Group 1 salts dissolve in water to give colourless solutions. Transition metal compounds are coloured.

Table 6.11 Test for sulfate ions

Ion	Sulfate
Result on adding barium chloride solution	White ppt of barium sulfate 
Ionic equation	$Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$

When solutions of potassium sulfate and barium chloride are mixed, a white precipitate of barium sulfate is formed as the positive barium ions from the barium chloride solution react with the negative sulfate ions of the potassium sulfate.

Word equation:
potassium sulfate + barium chloride → barium sulfate + potassium chloride

Symbol equation: $K_2SO_4(aq) + BaCl_2(aq) \rightarrow BaSO_4(s) + 2KCl(aq)$

Ionic equation: $Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$

3 Test for carbonate ions

Method

- Add dilute hydrochloric acid to the solid sample.
- Test the gas produced with limewater.
- If the limewater changes from colourless to milky then a carbonate is present.

Table 6.12 Test for carbonate ions

Ion	Carbonate
Result on adding dilute hydrochloric acid and testing the gas with limewater	Bubbles; the gas changes colourless limewater to milky 
Ionic equation	$CO_3^{2-}(s) + 2H^+(aq) \rightarrow CO_2(g) + H_2O(l)$

In this reaction, the positive hydrogen ions of the acid react with the negative carbonate ions to form carbon dioxide and water. Bubbles of carbon dioxide are observed and the identity of the gas is confirmed by testing with limewater.

Word equation:
calcium carbonate + hydrochloric acid → calcium chloride + water + carbon dioxide

Symbol equation: $CaCO_3(aq) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$

Ionic equation: $2H^+(aq) + CO_3^{2-}(aq) \rightarrow CO_2(g) + H_2O(l)$

This is also the test for a hydrogencarbonate. The ionic equation is

$$H^+(aq) + HCO_3^-(aq) \rightarrow CO_2(g) + H_2O(l)$$

H

F

Test yourself

24 Write down the formula of the anion present based on the results of the test given.

a) A white ppt is produced on adding a few drops of silver nitrate solution to a solution of the salt.

b) A white ppt is produced on adding a few drops of barium sulfate solution to a solution of the salt.

c) A cream ppt is produced on adding a few drops of silver nitrate solution to a solution of the salt.

d) On adding dilute nitric acid to the salt, a gas is produced which changes colourless limewater
to milky.

25 Write ionic equations, with state symbols, for the precipitation reactions of the following solutions.

a) barium chloride and copper sulfate

b) silver nitrate and potassium chloride

c) silver nitrate and sodium iodide

d) sodium sulfate and barium chloride

e) sulfuric acid and barium chloride

f) hydrochloric acid and silver nitrate.

26 Write balanced symbol equations, with state symbols, for the reactions in Question 25.

27 Write down the colour of each precipitate produced in Question 25.

28 Test tube A contains dilute nitric acid and test tube B contains dilute hydrochloric acid. Which one of the following substances, when added to each test tube, will produce a different observation in each?

A barium chloride solution

B copper(II) carbonate

C silver nitrate solution

D sodium hydroxide solution

Show you can ?

In an experiment, solutions containing each of barium, cadmium and silver ions were added separately to solutions containing carbonate, chloride or sulfate ions. The table shows which anions formed precipitates with the three different cations. A tick (✓) indicates that a precipitate formed and a cross (×) indicates that no precipitate formed.

	Anions		
Metal cations	Carbonate	Chloride	Sulfate
Barium	✓	×	✓
Cadmium	✓	×	×
Silver	✓	✓	✓

A student has a solution containing a mixture of barium, cadmium and silver ions. He wishes to make a solid precipitate containing each metal ion from the mixture and separate it by filtering. In which order must he add the anion solutions to enable him to filter out a compound of each metal in turn?

Planning experiments to identify cations and anions present in an unknown or a given compound

Chemists can use a variety of cation and anion tests to identify compounds. Table 6.13 is a useful summary of all the tests you need to enable you to plan the identification of a compound.

Table 6.13 A summary of some common tests for ions

Ion	Test	Result
Lithium	Flame test	Crimson flame
Sodium	Flame test	Yellow/orange flame
Potassium	Flame test	Lilac flame
Calcium	Flame test	Brick-red flame
Copper(II)	Flame test or Dissolve: add sodium hydroxide solution until in excess Dissolve: add ammonia solution until in excess	Blue-green/green-blue flame Blue ppt which remains on adding excess sodium hydroxide solution Blue ppt which dissolves in excess ammonia solution to give a deep blue solution
Iron(II)	Dissolve: add sodium hydroxide or ammonia solution	Green ppt
Iron(III)	Dissolve: add sodium hydroxide or ammonia solution	Red-brown ppt
Magnesium	Dissolve: add sodium hydroxide or ammonia solution until in excess	White ppt which remains on adding excess sodium hydroxide/ammonia solution
Aluminium	Dissolve: add sodium hydroxide or ammonia solution	White ppt which remains on adding excess sodium hydroxide solution but dissolves in excess ammonia solution
Zinc	Dissolve: add sodium hydroxide or ammonia solution	White ppt which dissolves in excess ammonia
Chloride	Dissolve then add silver nitrate solution	White ppt
Bromide	Dissolve then add silver nitrate solution	Cream ppt
Iodide	Dissolve then add silver nitrate solution	Yellow ppt
Sulfate	Dissolve then add barium chloride	White ppt
Carbonate	Add dilute acid to the solid and test gas with limewater	Bubbles Colourless limewater changes to milky

Tip

It is good practice to write sodium hydroxide solution rather than 'sodium hydroxide'.

Example

Plan an experiment to positively identify both ions in a sample of copper(II) sulfate.

Answer

Start with the positive ion, copper(II). There are three different ways of identifying copper ions and any one of them is suitable. In your answer you need to give an experimental method and also the expected result for a positive test.

- Carry out a flame test by dipping a nichrome wire loop into concentrated hydrochloric acid and then into the sample and holding it in a blue Bunsen burner flame. A blue-green/green-blue flame indicates the presence of copper(II) ions.

Or

- Dissolve some of the solid in water and add a few drops of sodium hydroxide solution, followed by excess sodium hydroxide solution. A blue ppt, which remains on adding excess, indicates copper ions.

Or

- Dissolve some of the solid in water and add a few drops of ammonia solution followed by excess ammonia solution. A blue ppt, which dissolves in excess ammonia to give a blue solution, indicates the presence of copper(II) ions.

Now look at the negative ion, sulfate, and describe the method of the test and the result for a positive test.

- Dissolve some of the solid in water and add a few drops of barium chloride solution. A white ppt indicates the presence of sulfate ions.

Prescribed practical

Practical C4: Identify the ions in an ionic compound using chemical tests

A student carried out a series of tests on a mixture of two ionic compounds labelled Z. The compounds have a common cation.

1 Copy and complete the table by writing an appropriate deduction for each observation.

Test	Observation	Deduction
Make a solution of Z by dissolving a spatula of Z in a test tube half full of deionised water. Warm gently.	White solid dissolves and colourless solution formed	
Add a few drops of sodium hydroxide to the solution prepared in Test 1, then add a further 3 cm^3 of the sodium hydroxide solution.	A white ppt forms which is insoluble in excess sodium hydroxide.	
Make a solution of Z by dissolving half a spatula of Z in a test tube half full of nitric acid solution.	No effervescence	
Transfer 1 cm^3 of the solution from Test 3 into each of two separate test tubes. Add a few drops of silver nitrate solution to the first test tube. Add a few drops of barium chloride solution to the second test tube.	 A white ppt forms A white ppt forms	

2 Name the two salts present in Z.

Enhance practical and enquiry skills with full coverage of the CCEA prescribed practicals

Test yourself

29 A green ppt was produced when sodium hydroxide solution was added to a solution of compound C. A yellow precipitate was produced when some silver nitrate solution was added to a separate solution of compound C.

a) Name the cation in C.
b) Name the anion in C.
c) Write the name and formula for C.
d) Write an ionic equation, including state symbols, for the formation of the green ppt.
e) Write an ionic equation for the formation of the yellow ppt.

30 Compound D was analysed and the following results were observed. It gave a crimson flame in a flame test. When dilute acid was added to a sample of D, a gas was given off that changed limewater from colourless to milky.

a) Name the positive ion in compound D.
b) Name the negative ion in compound D.
c) Write the name and formula of compound D.
d) Write an ionic equation for the reaction of D with dilute acid.

31 Some tests were carried out on a compound X.

Test	Observation
1 Describe X	Blue-green crystals
2 Make a solution of X by dissolving a spatula of X in a boiling tube of deionised water. Divide this solution into three test tubes for use in Tests 3, 4 and 5.	Blue solution formed
3 Add 1 cm^3 of nitric acid followed by 1 cm^3 of silver nitrate solution	No effervescence White ppt formed
4 Add 1 cm^3 of barium chloride solution	No reaction
5 Carry out a flame test on the compound	Blue-green flame
6 a) Add three drops of sodium hydroxide solution b) Add excess sodium hydroxide solution to the same test tube.	a) Blue ppt formed b) Blue ppt remains

a) Give an appropriate deduction for Tests 3, 4, 5 and 6.
b) Name and give the formula of compound X.
c) Give the formula of the white precipitate formed in Test 3.
d) Give the formula of the blue precipitate formed in Test 6.

Practice questions

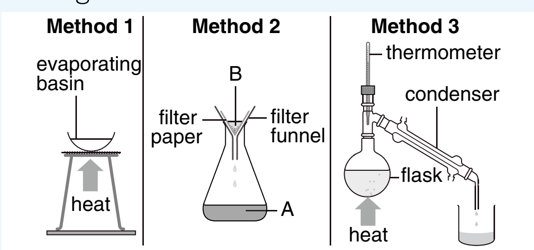
1 **a)** Some soluble aspirin tablets are dissolved in potable water to form a solution.

- **i)** What is potable water? *(1 mark)*
- **ii)** What is a solution? *(1 mark)*
- **iii)** Describe three steps needed to make groundwater potable. *(3 marks)*

b) Aspirin tablets contain the active drug aspirin, a binding agent and a sweetener.

- **i)** Why is an aspirin tablet considered to be a formulation? *(1 mark)*
- **ii)** How would you prove that a sample of the active drug was pure? *(1 mark)*
- **iii)** Apart from drugs, state two other examples of formulations. *(2 marks)*

2 Mixtures may be separated in the laboratory in many different ways. Three different methods of separating mixtures are shown below.



a) Name each method of separation. *(3 marks)*

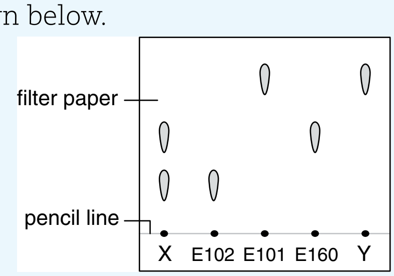
b) Which method (1, 2 or 3) would be most suitable for obtaining water from potassium chloride solution? *(1 mark)*

c) Which method would be most suitable for removing sand from a mixture of sand and water? *(1 mark)*

d) What general term is used for liquid A and solid B in Method 2? *(2 marks)*

e) State why Method 2 would not be suitable to separate copper(II) sulfate from copper(II) sulfate solution. *(1 mark)*

3 To determine whether two different orange drinks, X and Y, contained the food colourings E102, E101 or E160, a student put a drop of each orange drink and a drop of each food colouring along a pencil line on a piece of filter paper. The filter paper was placed in a solvent and the coloured components separated out. The results are shown below.



a) What is the name of the process used by the student to analyse the two orange drinks? *(1 mark)*

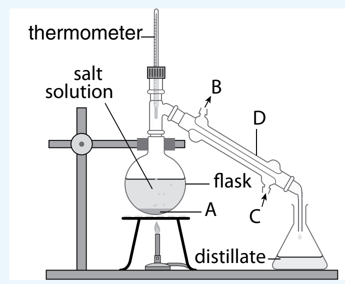
b) Orange drink X contains the food colouring E102. How do the results show this? *(1 mark)*

c) What other food colouring does orange drink X contain? *(1 mark)*

d) Re-draw the diagram and add a spot to show that orange drink Y contained food colouring E160. *(1 mark)*

e) The line across the bottom of the filter paper was drawn with a pencil not with ink. Why should the line not be drawn with ink? *(1 mark)*

4 Drinking water has been distilled from sea water since at least 200AD when the process was described clearly by the Greek philosopher Aristotle. The distillation apparatus used to separate pure water from salt solution in the laboratory is shown below.



a) Name the solute in salt solution. *(1 mark)*

b) Name the solvent in salt solution. *(1 mark)*

c) What is solid A? Suggest its purpose. *(2 marks)*

d) What labels should be inserted at positions B and C? *(2 marks)*

e) Name the piece of apparatus labelled D. *(1 mark)*

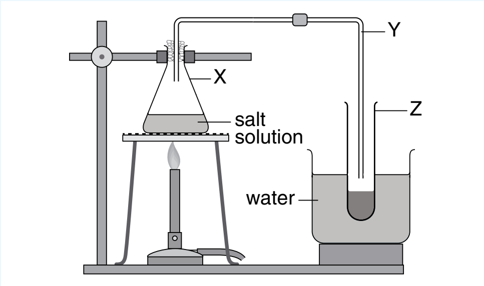
f) What is the temperature on the thermometer when distillation is occurring? What does this suggest? *(2 marks)*

g) What happens to the salt dissolved in the solution during this process? *(1 mark)*

h) Distillation involves evaporation. What is meant by evaporation? Where in the apparatus does distillation take place? *(2 marks)*

i) Distillation involves condensation. What is meant by condensation? Where in the apparatus does condensation take place? *(2 marks)*

j) What advantages does the apparatus shown in Diagram 1 have over the apparatus in Diagram 2? *(1 mark)*



k) What is the purpose of the beaker of water in Diagram 2? *(1 mark)*

l) Name the pieces of apparatus X, Y and Z. *(3 marks)*

m) Describe how you would test the purity of the distillate. *(1 mark)*

n) This method is used as a desalination process. What is desalination? *(1 mark)*

o) How would you test and prove that the distillate is water? *(2 marks)*

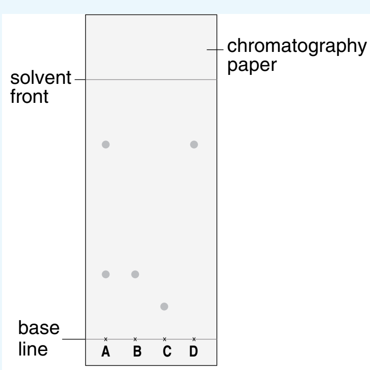
5 To determine whether a sample of water contained some dissolved metal ions, a chromatography experiment was carried out using the water sample (A) and known metal ion solutions B (containing copper(II) ions), C (containing iron(II) ions) and D (containing iron(III) ions). The method used was:

- Draw a base line on the chromatography paper 1.5 cm from the bottom using a pencil.
- Place a concentrated drop of each solution to be tested on the base line.
- Place the chromatography paper into a chromatography tank containing water at a depth of 1 cm.
- After the water had soaked up the paper, it was dried and sprayed with sodium hydroxide solution and coloured spots appeared on the paper.

a) Why is it necessary to use a pencil line rather than a line drawn with a pen? *(1 mark)*

b) Why is the water solvent at a depth of only 1 cm? *(1 mark)*

c) How is a concentrated drop of each solution added to the chromatography paper? *(2 marks)*



The chromatogram obtained is shown below.

d) Name two metal ions which are present in the water sample A. *(2 marks)*

e) Write the formula of the compound formed when spot C reacts with sodium hydroxide. *(1 mark)*

f) Write an equation which is used to calculate R_f value and calculate the R_f value for spot B. *(2 marks)*

g) If the experiment was repeated using a different solvent, would the R_f value be the same? *(1 mark)*

6 Compound S is an ionic salt. To identify the positive and negative ions present, some tests were carried out. A white precipitate was produced when sodium hydroxide solution was added to a solution of S. When barium chloride solution was added to a separate solution of S, a white precipitate was also produced.

a) Identify the negative ion in compound S. *(1 mark)*

b) Comment on the validity of the test for the positive ion and suggest any ways in which validity could be improved. *(3 marks)*

c) Write an ionic equation, including state symbols, for the formation of the white precipitate when barium chloride is added. *(3 marks)*

7 a) In an experiment to determine which Group 7 ion was present in each of three different acids, a few drops of silver nitrate solution were added to samples of the acid solutions.

i) Copy and complete the table below to show the results of these tests.

Acid	**Observation on addition of a few drops of aqueous silver nitrate**
Hydrobromic acid	
Hydrochloric acid	
Hydroiodic acid	

ii) Write a ionic equation for the reaction of hydrochloric acid with silver nitrate solution. *(2 marks)*

b) Hydrochloric acid reacts with bases to form salts such as sodium chloride and magnesium chloride. An antiseptic mouthwash is thought to contain both of these salts.

i) Describe how you would confirm that the mouthwash contained sodium ions. *(2 marks)*

ii) Describe how you would confirm experimentally that the mouthwash contained magnesium ions. In your answer, refer to the validity of your test. *(5 marks)*

8 A mixture of two ionic compounds was analysed to determine the ions present in the mixture. The two ionic compounds have the same anion. The results of the tests are given in the table below. Use the information in the table to answer the questions that follow.

Description of test	Observations
Test 1 A flame test was carried out on a solid sample of the mixture	Lilac flame
Test 2 A sample of the mixture was dissolved in deionised water and divided into two test tubes. Five drops of sodium hydroxide solution was added to one test tube, followed by excess. Five drops of ammonia solution was added to the other test tube, followed by excess. White precipitate which dissolves in excess ammonia solution	White precipitate which dissolves in excess sodium hydroxide solution
Test 3 A sample of the mixture was dissolved in deionised water and barium chloride solution was added.	White precipitate

a) Using the evidence from Test 1 only, name a cation which is present in the mixture. *(1 mark)*

b) Using the evidence from Test 2 only, name a cation which may be present in the mixture. *(1 mark)*

c) Using the evidence from Test 3, write the formula of the anion which is present in the mixture. *(1 mark)*

d) Suggest the names of two compounds which may be present in the mixture. *(2 marks)*

e) Describe how you would carry out Test 1 experimentally. *(3 marks)*

9 a) The presence of anions in salts can be determined by precipitation reactions.

i) What is an anion? *(1 mark)*

ii) What is meant by the term precipitate? *(1 mark)*

iii) Name the solution which is used to test for the presence of sulfate ions. *(1 mark)*

iv) Potassium iodide solution was mixed with silver nitrate solution and a precipitate formed. State the colour of the precipitate. *(1 mark)*

b) i) The presence of cations in salts can be determined using a flame test or by adding sodium hydroxide solution or ammonia solution. Using your knowledge of the results of such tests, copy and complete the table below. *(5 marks)*

Metal ion	Copper ion	Aluminium ion
Flame test result		
Result on adding a few drops of sodium hydroxide solution, followed by excess, to the metal ion solution		
Result on adding a few drops of ammonia solution, followed by excess, to the metal ion solution		

c) The presence of iron(II) or iron(III) ions in an iron supplement may be detected by adding sodium hydroxide solution to a solution of the supplement.

i) State the observation if iron(II) ions were present in the supplement. *(1 mark)*

ii) Write an ionic equation, including state symbols, for the reaction of iron(II) ions in solution with sodium hydroxide solution. *(4 marks)*

iii) State the observation if iron(III) ions were present in the supplement. *(1 mark)*

iv) Name another reagent which could be used, instead of sodium hydroxide solution, to test for the presence of iron(III) ions. *(1 mark)*

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