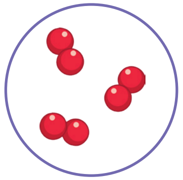
AQA GCSE Chemistry (Separate Science) Unit 8: Chemical Analysis

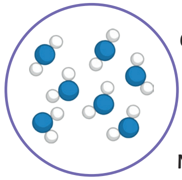
Pure Substances

Pure substances, in chemistry, only contain **one type of element** or **one type of compound**. For example, pure water will just contain water (a compound).

In our everyday language, we use the word 'pure' differently to how it is used in chemistry. Pure can mean a **substance** that has had **nothing else added to it** and is in its natural state. An example of this is pure orange juice. This means that the bottle will just contain orange juice and no other substances.

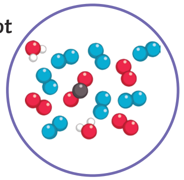
Elements are made up of **one type of atom**. For example, oxygen is made up of oxygen atoms. Carbon is made up of carbon atoms.





Compounds are **two or more elements** that are **chemically joined** together. For example, NaCl which is sodium chloride.

Mixtures are **two or more elements or compounds** that are **not chemically joined** together. An example of this is a standard cup of coffee. Coffee contains water, milk, coffee and possibly sugar. The components of the cup of coffee are not bonded together.



Pure Substances have a **sharp melting point** compared to **impure** substances which **melt over a range** of temperatures.

Formulations

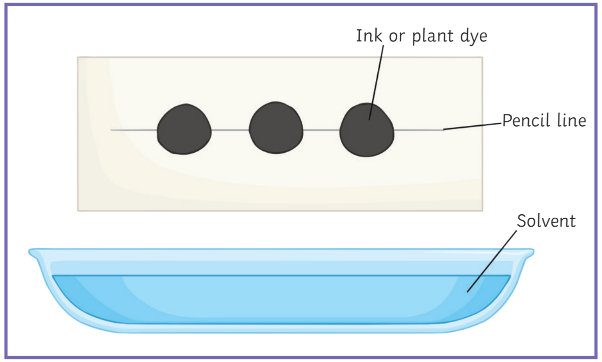
Formulations are **mixtures of compounds or substances** that **do not react together**. They do **produce a useful product** with desirable characteristics or properties to suit a particular function.

There are examples of formulations all around us such as medicines, cleaning products, deodorants, hair colouring, cosmetics and sun cream.

Chromatography

Paper chromatography is a separation technique that is used to **separate** mixtures of **soluble substances**. How soluble a substance is determines how far it will travel across the paper.

In chromatography, there are **two phases**: the mobile and **stationary** phase.

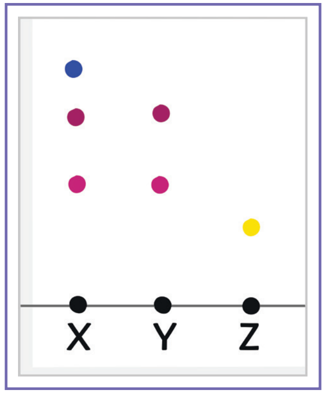


The **mobile phase moves** through the stationary phase. The **solvent** is the **mobile phase**. It moves through the paper carrying the different substances with it.

The **stationary phase** in paper chromatography is the **absorbent paper**.

Separation of the dissolved substances produces what is called **chromatogram**. In paper chromatography, this can be used to **distinguish** between those substances that are **pure** and those that are **impure**. **Pure substances** have **one spot** on a chromatogram as they are made from a single substance. **Impure substances** produce **two or more spots** as they contain multiple substances.

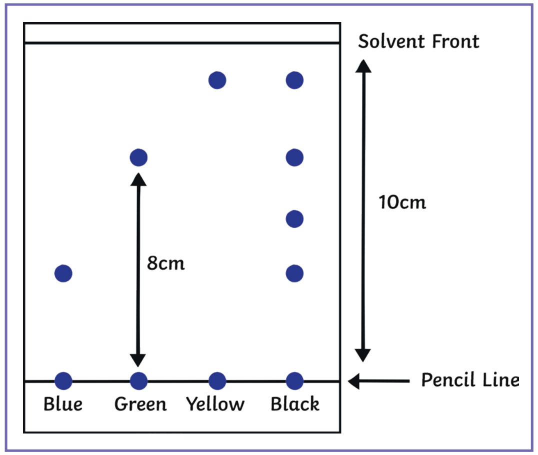
By calculating the **R_f values** for each of the spots, it is possible to identify the unknown substances. Similarly, if an unknown substance produces the **same number** and **colour of spots**, it is possible to match it to a known substance.



R_f Value

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

Different compounds have different R_f values in different solvents. The R_f values of known compounds can be used to help identify unknown compounds.



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Required Practical – Paper Chromatography

Investigate how paper chromatography can be used to separate and distinguish between coloured substances.

Step 1 – Using a ruler, measure 1cm from the bottom of the chromatography paper and mark with a small dot using a pencil. Rule a line across the bottom of the chromatography paper with a pencil, going through the dot you have just made.

Step 2 – Using a pipette, drop small spots of each of the inks onto the pencil line. Leave a sufficient gap between each ink spot so that they do not merge.

Step 3 – Get a container and pour a suitable solvent into the bottom. The solvent should just touch the chromatography paper. The solvent line must not go over the ink spots as this will cause the inks to run into each other.

Step 4 – Place the chromatography paper into the container and allow the solvent to move up through the paper.

Step 5 – Just before the solvent line reaches the top of the paper, remove the chromatogram from the container and allow to dry.

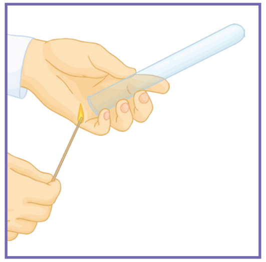
Step 6 – Once the chromatogram has dried, measure the distance travelled by the solvent.

Step 7 – Measure the distance travelled by each ink spot.

Step 8 – Calculate the R_f value. Compare the R_f value for each of the spots of ink.

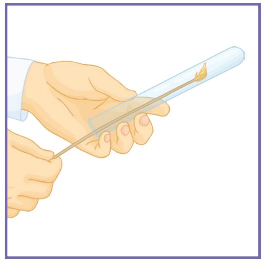
$$R_f = \frac{\text{distance travelled by substance}}{\text{distance travelled by solvent}}$$

Identification of the Common Gases



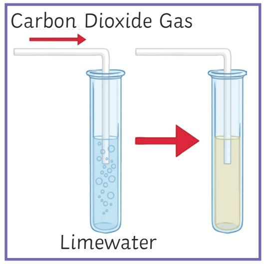
The Test for Hydrogen

Place a burning splint at the opening of a test tube.
If hydrogen gas is present, it will burn rapidly with a **squeaky-pop sound**.



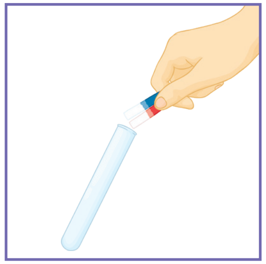
The Test for Oxygen

Place a glowing splint inside a test tube.
The **splint will relight** in the presence of oxygen.



The Test for Carbon Dioxide

Calcium hydroxide (lime water) is used to test for the presence of carbon dioxide. When carbon dioxide is bubbled through or shaken with limewater, the limewater turns **cloudy**.



The Test for Chlorine

Damp litmus paper is used to test for chlorine gas.
The litmus paper becomes **bleached and turns white**.

Flame Tests

Metal ions when heated produce a variety of flame colours. Flame tests are used to **identify** the **metal ion** that is present; each metal ion produces a different coloured flame.

Step 1 – Dip a wire loop into a sample of the solid compound being tested.

Step 2 – Place the loop into the flame of the Bunsen burner. Ensure that the Bunsen burner is set to a roaring blue flame.

Step 3 – Observe the colour of the flame produced and record it in a table.

Mixtures of ions may cause some flame colours to not be as clear.

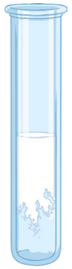
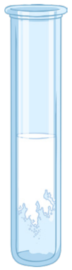
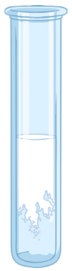
Ion	Colour of the Flame
Li^{+}	 crimson
Na^{+}	 yellow
K^{+}	 lilac
Ca^{2+}	 orange-red
Cu^{2+}	 green

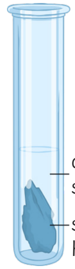
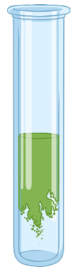
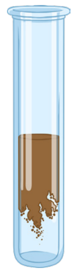
Metal Hydroxides

In order to identify metal ions, **sodium hydroxide solution** is added. Solutions of calcium, magnesium and aluminium all form white precipitates. Only the aluminium hydroxide **precipitate** dissolves in excess sodium hydroxide. Iron (II), iron (III) and copper (II) all form coloured precipitates when sodium hydroxide solution is added.

magnesium sulfate + sodium hydroxide ⟶ magnesium hydroxide + sodium sulfate

$MgSO_4 + 2NaOH \longrightarrow Mg(OH)_2 + Na_2SO_4$

Ion	Colour of the Precipitate Produced
Al^{3+}	white 
Ca^{2+}	white 
Mg^{2+}	white 

Ion	Colour of the Precipitate Produced
Cu^{2+}	blue 
Fe^{2+}	green 
Fe^{3+}	brown 

Ionic Equations

An ionic equation can be used to represent each of the **precipitation** reactions. These equations only show the ions that are involved in the precipitation reaction. The equations do not show the sodium or sulfate ions. This is because these ions are called spectator ions. **Spectator ions** are ions that do not take part in the chemical reaction.

Copper (II)

$Cu^{2+} + 2OH^- \longrightarrow Cu(OH)_2$

Copper has **lost** two **negative charges**, hence why copper is Cu^{2+}. In order to balance out this loss of charges, the copper ion **must gain** two negative charges. These negative charges come in the form of **two OH^- ions**.

Iron (III)

$Fe^{3+} + 3OH^- \ Fe(OH)_3$

Iron (III) has **lost** three **negative charges**, hence why iron is Fe^{3+}. In order to balance out the loss of charges, the iron ion **must gain** three negative charges. These negative charges come in the form of **three OH^- ions**.

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Testing for Carbonate Ions (CO_3^{2-}) Chemistry Only

Place a small volume of limewater into a test tube. In a separate test tube, add a small sample of the **carbonate** and add a few drops of **hydrochloric acid** (acids are a source of H^+ ions) using a pipette. Seal the test tube with a bung connected to a delivery tube; the delivery tube should be placed in the test tube containing the limewater. Bubbles of **carbon dioxide** gas will be produced. The **limewater will turn a milky colour** indicating a positive test for carbon dioxide.

$CO_3^{2-} + 2H^+ \longrightarrow CO_2 + H_2O$

Testing for Sulfate Ions (SO_4^{2-})

Using a pipette, add a few drops of **barium chloride** solution to the sample followed by a few drops of **hydrochloric acid**. A positive result for sulfate ions will produce a white precipitate.

$SO_4^{2-} + Ba^{2+} \longrightarrow BaSO_4$

Testing for Halide Ions (I^-, Br^-, Cl^-)

Using a pipette, add a few drops of dilute **nitric acid** to the sample followed by a few drops of **silver nitrate solution**. Leave it to stand and **observe the colour** of the **precipitate formed**.

Each halide ion produces a different coloured precipitate.

- **Chloride** produces a **white** precipitate.
- **Bromide** ions produce a **cream** precipitate.
- **Iodide** ions produce a **yellow** precipitate.

Flame Emission Spectroscopy

Flame emission spectroscopy is an instrumental method of analysis. The benefits of instrumental methods of analysis are that it is **rapid, accurate and sensitive**. The drawbacks to such methods are that the equipment is often **expensive** and **requires special training** to use.

Flame emission spectroscopy is a technique that is used to **identify** metal ions in solution. The samples that are tested normally include biological fluids and tissues.

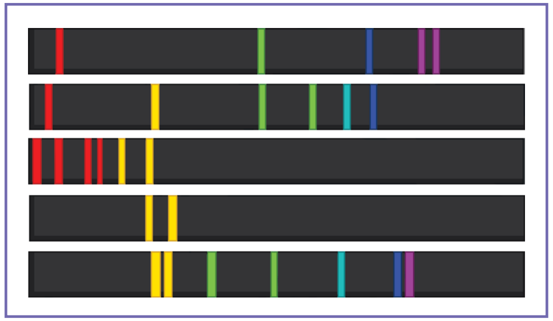
How It Works

Step 1 – A sample is heated in a flame.

Step 2 – Electrons in the metal ions are excited by the thermal energy provided from the flame. As a result, the electrons move into a higher energy level.

Step 3 – When the electrons fall back into a lower energy level, they release energy in the form of light.

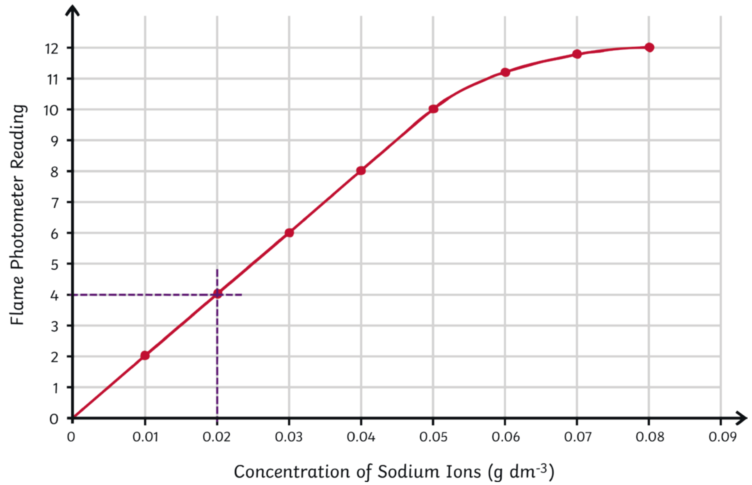
Step 4 – The emitted wavelengths of light are analysed instrumentally.



Step 5 – To identify the metal present, its spectrum is compared with reference spectra from known metal ions.

Above is an example of the spectra produced by flame emission spectroscopy. It looks like a colourful array of lines. **Each metal ion produces a unique emission spectrum**.

Calibration Curve



The readings for different concentrations of metal ions in solutions are taken. These readings are then used to plot a calibration curve.