A **pure substance** is a simgle element or compound.

A **mixture** is made up of two or more different elements or compounds that are not chemically joined together.

A **formulation** is a mixture that has been specifically designed to produce a useful product. Examples include paints and medicinal drugs.

Identifying pure substances

The melting and boiling point of a substance can be used to tell whether it is pure or not.

There are two ways you can use the melting and boiling point to tell if the sample is pure.

1) The **range** of the melting and boiling points will be very small if the sample is pure. Example:

	Pure caffeine	Impure caffeine
Melting point/ °C	234-237	180-220

2) Alternatively, once you have recorded the melting/boiling point for a substance you can **compare** it to a **database** of known values.

- If the melting point you recorded matches the melting point given in the database, it is pure.
- If the melting point does not match the value in the database, your sample is impure.
- Generally impure substances will have a lower melting point and a higher boiling point than the pure substance.

Chromatography

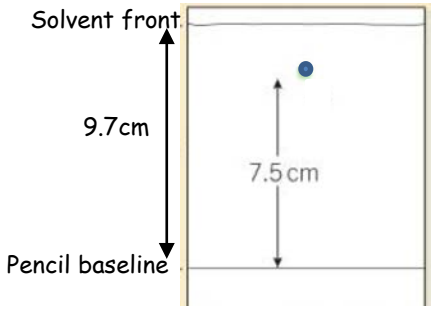
Chromatography is used to identify unknown compounds in a sample.

- The **mobile phase** moves and carries the different compounds in the sample through the **stationary phase**.
- In **paper chromatography**, the mobile phase is a solvent such as water which carries the sample along the paper (stationary phase).
- The **different compounds** in the sample will **travel different distances** along the paper.
- If a compound is more **strongly attracted** to the mobile phase (very soluble in it) than the stationary phase, it will **travel further** up the paper.

Pure or a mixture? Read the chromatogram vertically. If there is **only** one spot above the point where the sample was placed the sample is **pure**, more than one spot means a mixture.

Calculating the Retention Factor value:

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



Example

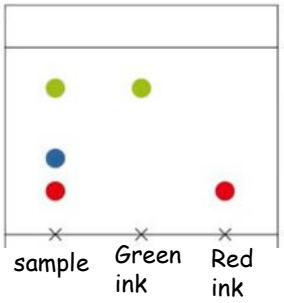
$$\frac{7.5\text{cm}}{9.7\text{cm}} = 0.77$$

Identifying unknown compounds

To identify the unknown compounds we calculate the Rf value (see left).
When we have calculated the Rf value we can **compare** that to **Rf values** of known substances which are **stored in databases**.

Alternatively, when we carry out the chromatography we can also run known samples and see if any spots travel the same distance.

Example:



In this example, two known compounds (red and green ink) have been added to the chromatography paper. We can use them to confirm that the unknown sample contains red and green ink because the spots in the sample have travelled the same vertical distance.

- The baseline is **drawn in pencil** and not ink because ink will dissolve in water and run.
- The pencil baseline must be **above** the surface of the water otherwise the sample will dissolve into the beaker of water and be lost.

Testing for gases

Hydrogen – Hold a lit splint near the sample of gas.
Positive test: Hydrogen will burn with a squeaky pop sound.

Carbon dioxide – Bubble gas through limewater
Positive test: If carbon dioxide is present it will turn limewater cloudy.

Oxygen – Hold a glowing splint near the sample of gas.
Positive test: If oxygen is present a glowing splint will relight.

Chlorine – Use damp blue litmus paper
Positive test: Chlorine will bleach damp blue litmus paper white.

Flame tests for metal ions

Dip a nichrome wire loop into concentrated sulphuric acid and then heat it to clean it.

Then, dip the wire in the unknown sample before holding it in a roaring blue flame of a Bunsen burner.

The colour of the flame will tell you which metal ion is present. For mixtures, remember certain colours like the yellow of sodium may mask other colours that are present.

Metal ion	Flame colour
Lithium Li^+	Crimson
Sodium Na^+	Yellow
Potassium K^+	Lilac
Calcium Ca^{2+}	Orange-red
Copper Cu^{2+}	Green

Metal cation tests with sodium hydroxide

Sodium hydroxide can be used to identify positive metal ions.

The following ions form a **white precipitate** with **sodium hydroxide** solution:

aluminium Al^{3+}
calcium Ca^{2+}
magnesium Mg^{2+}

The ions react with sodium hydroxide to produce insoluble metal hydroxides for example:

Ionic equation

$$Al^{3+}(aq) + 3OH^-(aq) \rightarrow Al(OH)_3(s)$$

Note: To help tell the 3 ions apart if you continue to add sodium hydroxide the **aluminium precipitate will dissolve**. You can also use a flame test.

Metal cation tests with sodium hydroxide continued

Some metal ions form coloured precipitates when **sodium hydroxide** is added

Copper (II) Cu^{2+} forms a **blue** precipitate
Iron (II) Fe^{2+} forms a **green** precipitate
Iron (III) Fe^{3+} forms a **brown** precipitate

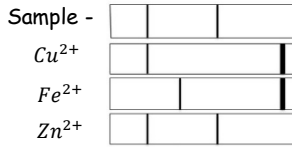
Instrumental analysis

Many industries need quick and accurate methods for analysing their products. Modern equipment is very **accurate**, **quick** and can work on **small samples** although it is **expensive**.

An example of instrumental analysis is flame emission spectroscopy. Metal ions emit unique wavelengths of light when they are heated.

In a spectrometer we can analyse the unique pattern of radiation (a line spectrum) and compare it to other know line spectrums.

In this example we can confirm the sample contains Zn^{2+} ions because the line spectrums match.



Carbonates

If a metal carbonate is present, it will react with acid and release carbon dioxide meaning fizzing is observed.

Metal carbonate + acid → salt + water + carbon dioxide

The reaction can be shown by an ionic equation;

$$CO_3^{2-}(s) + 2H^+(aq) \rightarrow CO_2(g) + H_2O(g)$$

Carbonate ion — acid

The presence of carbon dioxide can be confirmed using limewater.

Halides (chlorides, bromides and iodides)

To test for halides add **dilute nitric acid** and then **silver nitrate** solution. If a precipitate forms, a halide is present.

The nitric acid is added to remove any carbonate ions because they form precipitates with silver ions.

chloride ions: **white** precipitate
bromide ions: **cream** precipitate
iodide ions: **yellow** precipitate

Example

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

Ionic equation (where X represents a halide ion)

$$Ag^+(aq) + X^-(aq) \rightarrow AgX\,(s)$$

Sulphates

You can test for sulphate ions by adding dilute **hydrochloric acid** and **barium chloride** solution. If a **white precipitate** forms, sulphate ions are present.

The hydrochloric acid is added to remove any carbonate ions which would form a precipitate with the barium ions.

Example: If the unknown compound contained potassium sulphate, the equation would be:

$$K_2SO_4\,(aq) + BaCl_2\,(aq) \rightarrow 2KCl(aq) + BaSO_4(s)$$

Ionic equation

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