

**Section 1: Key terms**

Pure	A pure substance is a single <b>element</b> or <b>compound, not mixed</b> with anything else.
Formulation	<b>Useful mixtures</b> that have a precise purpose. The quantity of each component in a formulation has been measured carefully. Formulations include <b>fuels, cleaning agents, paints, medicines, alloys, fertilisers &amp; foods.</b>
Melting point	The <b>temperature</b> at which a substance turns from a solid to a liquid.
Boiling point	The <b>temperature</b> at which a substance turns from a liquid to a gas.
Chromatography	An analytical method used to <b>separate substances</b> in a <b>mixture.</b>
R <sub>f</sub> value	<b>Retention factor.</b> A ratio, calculated by dividing the distance moved by a spot up the paper by the distance the solvent front travels.
Solvent	The chemical that <b>dissolves the sample</b> in chromatography.
Solvent front	The <b>maximum distance</b> the <b>solvent moves</b> up the paper.
Stationary phase	The phase where the molecules can't move. The <b>medium</b> (e.g. paper) through which the <b>mobile phase passes</b> in <b>chromatography.</b>
Mobile phase	The phase (in chromatography), where <b>molecules can move.</b> The <b>solvent</b> (e.g. water) that carries the sample (e.g. ink).
Energy levels	Electrons orbit the nucleus in specific energy levels (or <b>shells</b> ).

**Section 2: Pure substances and mixtures**

You can use melting points and boiling points to identify pure substances. The **test for pure water** is that it **melts at exactly 0°C** and **boils at exactly 100°C.** These fixed points can be looked up in data books.

A **mixture does not** have a **sharp melting point or boiling point**, it changes state over a **range of temperatures.**

**Impurities** will **lower the melting point** of a substance and **increase its boiling point.** The purer the compound is, the narrower the melting point range. Crude aspirin made in the lab has a melting point between 128-132°C, whereas pure aspirin has a sharp melting point of 136°C.

**Section 3: Formulations**

Formulations are important in the pharmaceutical industry and are made by mixing the components in **carefully measured quantities** to ensure that the product has the **required properties.**

When you buy a product, the ratio or percentage of each component is found on the packaging. This is its formulation.

Depending on the purpose of the product, the amount and type of chemicals used will be changed to make sure it is right for the job. E.g. Pigment of paint.

**Section 4: Paper Chromatography**

Chromatography is a physical method that is good for **separating and identifying** things. Chromatography always involves two phases, a **mobile phase** and a **stationary phase.**

In **paper chromatography,** the **mobile phase** is the **solvent,** the **stationary phase** is the **paper.**

During chromatography, the substances in the sample constantly move between the mobile and the stationary phase – an **equilibrium is formed between the two phases.**

A substance which has **stronger attraction** to the stationary phase will **not move very far up the paper** in the same time.

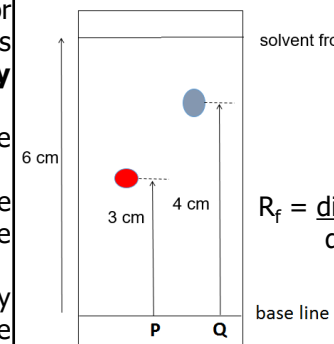
A substance which has **stronger attraction (solubility)** to the **mobile phase** will spend more time in the mobile phase and hence **move further up** the paper.

Different compounds have **different R<sub>f</sub>** values in **different solvents,** which can be used to help identify the compounds.

**Explaining how different dyes are separated using paper chromatography:** Solvent (mobile phase) moves through the paper (stationary phase). Different dyes have different solubilities in solvent and different attractions for the paper and hence are carried different distances.

**Section 5: Testing for gases**

Gas	Procedure	Positive result
Hydrogen	Hold a <b>lighted splint</b> at the end of a test tube producing the gas.	The lighted splint <b>"pops"</b> .
Oxygen	Hold a <b>glowing splint</b> in a test tube of the gas.	The glowing splint <b>"relights"</b> .
Carbon dioxide	Bubble gas through a solution of <b>limewater.</b>	The limewater turns <b>"milky"</b> .
Chlorine	When damp litmus paper is put into chlorine gas	The litmus paper is <b>"bleached"</b> and <b>turns white.</b>



The R<sub>f</sub> value can be calculated using the formula:

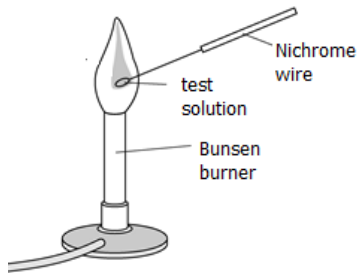
$$R_f = \frac{\text{distance moved by substance}}{\text{distance moved by solvent}}$$

$$\text{For substance P } R_f = \frac{3}{6} = 0.5$$

$$\text{For substance Q } R_f = \frac{4}{6} = 0.67$$

**Section 6a: Testing for positive ions : Flame tests**

To carry out a flame test: Dip a **nichrome wire** in **dilute hydrochloric acid**, then **dip it** into the **metal compound** that's being tested. Hold the loop of the wire in the **roaring blue flame** of the **Bunsen burner** and observe the flame colour.



Positive Ion	Flame colour
Lithium Li <sup>+</sup>	Crimson red
Sodium Na <sup>+</sup>	Yellow
Potassium K <sup>+</sup>	Lilac
Calcium Ca <sup>2+</sup>	Orange red
Copper Cu <sup>2+</sup>	Green
Magnesium Mg <sup>2+</sup>	No colour

**Section 8: Key terms**

Nichrome	An alloy of chromium and nickel.
Precipitate	An <b>insoluble</b> solid.
In excess	When you add something <b>in excess</b> , then you are adding more of it.
Effervesce	Bubbles/fizzing when a gas is produced.
Line spectra	An electromagnetic spectrum consisting of <b>discrete lines</b> , usually characteristic of excited atoms or molecules.
Ionic equation (HT)	An equation that shows only the ions or atoms that change in a reaction

**Section 6b: Testing for positive ions: with sodium hydroxide**

Dilute **sodium hydroxide** is added, **initially dropwise** and then until **in excess**.

Positive Ion	Test	Observation	Balanced equation examples (HT ionic equations in grey)
Aluminium Al <sup>3+</sup>	Add dilute sodium hydroxide	A <b>white precipitate</b> forms, the precipitate <b>dissolves</b> in <b>excess</b> sodium hydroxide solution.	Al <sub>2</sub> (SO <sub>4</sub> ) <sub>3</sub> (aq) + 6NaOH(aq) → Al(OH) <sub>3</sub> (s) + Na <sub>2</sub> SO <sub>4</sub> (aq) Al <sup>3+</sup> (aq) + 3OH <sup>-</sup> (aq) → Al(OH) <sub>3</sub> (s)
Calcium Ca <sup>2+</sup>	Add dilute sodium hydroxide	A <b>white precipitate</b> forms, the precipitate <b>does not dissolve</b> in <b>excess</b> sodium hydroxide solution.	CaSO <sub>4</sub> (aq) + 2NaOH(aq) → Ca(OH) <sub>2</sub> (s) + Na <sub>2</sub> SO <sub>4</sub> (aq) Ca <sup>2+</sup> (aq) + 2OH <sup>-</sup> (aq) → Ca(OH) <sub>2</sub> (s)
Magnesium Mg <sup>2+</sup>	Add dilute sodium hydroxide	A <b>white precipitate</b> forms, the precipitate <b>does not dissolve</b> in <b>excess</b> sodium hydroxide solution.	MgCl <sub>2</sub> (aq) + 2NaOH(aq) → Mg(OH) <sub>2</sub> (s) + 2NaCl(aq) Mg <sup>2+</sup> (aq) + 2OH <sup>-</sup> (aq) → Mg(OH) <sub>2</sub> (s)
Copper(II) Cu <sup>2+</sup>	Add dilute sodium hydroxide	A light blue <b>precipitate</b> forms	CuSO <sub>4</sub> (aq) + 2NaOH(aq) → Cu(OH) <sub>2</sub> (s) + Na <sub>2</sub> SO <sub>4</sub> (aq) Cu <sup>2+</sup> (aq) + 2OH <sup>-</sup> (aq) → Cu(OH) <sub>2</sub> (s)
Iron(II) Fe <sup>2+</sup>	Add dilute sodium hydroxide	A <b>green precipitate</b> forms, the precipitate slowly turns <b>brown</b> .	FeCl <sub>2</sub> (aq) + 2NaOH(aq) → Fe(OH) <sub>2</sub> (s) + 2NaCl(aq) Fe <sup>2+</sup> (aq) + 2OH <sup>-</sup> (aq) → Fe(OH) <sub>2</sub> (s)
Iron(III) Fe <sup>3+</sup>	Add dilute sodium hydroxide	A <b>reddish brown precipitate</b> forms.	FeCl <sub>3</sub> (aq) + 3NaOH(aq) → Fe(OH) <sub>3</sub> (s) + 3NaCl(aq) Fe <sup>3+</sup> (aq) + 3OH <sup>-</sup> (aq) → Fe(OH) <sub>3</sub> (s)

**Section 7: Testing for negative ions**

Negative Ion	Test	Observation	Balanced equation examples (HT ionic equations in grey)
Carbonate CO <sub>3</sub> <sup>2-</sup>	Add dilute acid	Solution <b>effervesces</b> (due to carbon dioxide being produced). The gas turns limewater milky.	MgCO <sub>3</sub> (s) + 2HCl(aq) → MgCl <sub>2</sub> (aq) + H <sub>2</sub> O(l) + CO <sub>2</sub> (g) CO <sub>3</sub> <sup>2-</sup> (s) + 2H <sup>+</sup> (aq) → H <sub>2</sub> O(l) + CO <sub>2</sub> (g)
Sulfate	Add dilute hydrochloric acid followed by barium chloride solution	A <b>white precipitate</b> forms	BaCl <sub>2</sub> (aq) + Na <sub>2</sub> SO <sub>4</sub> (aq) → BaSO <sub>4</sub> (s) + 2NaCl(aq) Ba <sup>2+</sup> (aq) + SO <sub>4</sub> <sup>2-</sup> (aq) → BaSO <sub>4</sub> (s)
Chloride	Add dilute nitric acid followed by silver nitrate solution	A <b>white precipitate</b> forms	AgNO <sub>3</sub> (aq) + NaCl(aq) → AgCl(s) + NaNO <sub>3</sub> (aq) Ag <sup>+</sup> (aq) + Cl <sup>-</sup> (aq) → AgCl(s)
Bromide	Add dilute nitric acid followed by silver nitrate solution	A <b>cream precipitate</b> forms	AgNO <sub>3</sub> (aq) + NaBr(aq) → AgBr(s) + NaNO <sub>3</sub> (aq) Ag <sup>+</sup> (aq) + Br <sup>-</sup> (aq) → AgBr(s)
Iodide	Add dilute nitric acid followed by silver nitrate solution	A <b>yellow precipitate</b> forms	AgNO <sub>3</sub> (aq) + NaI(aq) → AgI(s) + NaNO <sub>3</sub> (aq) Ag <sup>+</sup> (aq) + I <sup>-</sup> (aq) → AgI(s)

**Section 9: Instrumental analysis**

Many industries require fast and accurate methods for analysis of products and to test emissions produced during the manufacturing process. They use modern instrumental analysis (like **flame emission spectroscopy**) for this task.

**Advantages of modern instrumental analysis** (compared with traditional chemical tests). **Disadvantages of modern instrumental analysis** (compared with traditional chemical tests).

<p><b>Highly accurate</b> <b>Sensitive</b> <b>Faster</b> Only small amounts need to be tested Fewer people needed to carry out analysis</p>	<p>Expensive Required training to use Results have to be compared with data from known substances</p>
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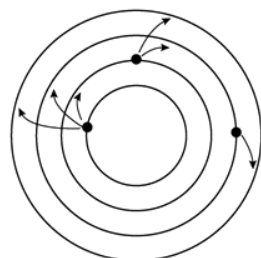
**Section 9a: Flame emission spectroscopy**

Flame emission spectroscopy is an instrumental method used to tell scientists which **metals ions** are present from their characteristic **line spectra** or the **concentration of metal ions** in a **solution**.

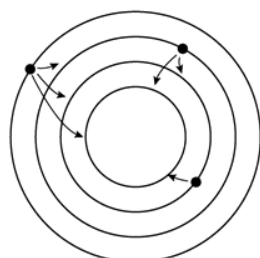
**Analysing metal ions from their line spectra**

**Theory** of flame tests (and flame emission spectroscopy)

- A sample is **heated in a blue flame**.
- The energy provided (from the flame) **excites electrons** in the metals ions making them **jump into higher energy levels** or shells (excited state).
- When they fall back down (relax) to lower energy levels (shells) energy is released as light energy.

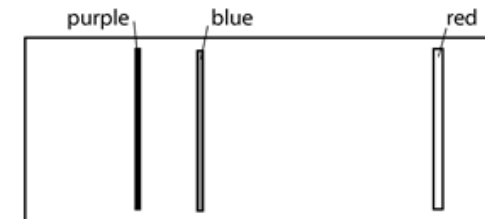


Excited electrons **jumping** into higher energy levels



Electrons **falling back down** into lower energy levels

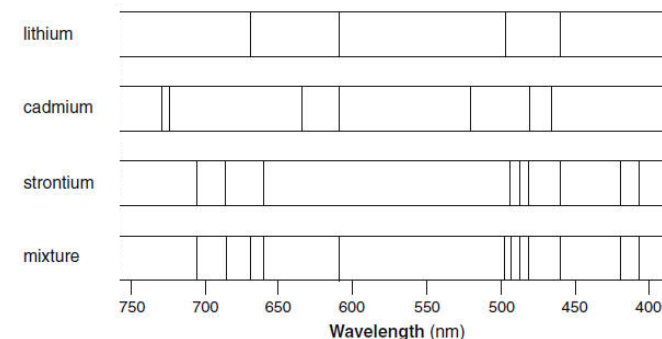
In the spectrometer, the wavelengths of light produced are analysed by passing it through a spectroscope. Each **metal ion** gives a **characteristic line spectrum** which can be used to **identify the metal** by comparing it with data from a **spectral data base**. Below is an example of a line spectrum of hydrogen.



The spectrum of hydrogen.

**Analysis**

Metal ions in a mixture can also be identified using a spectroscope. The spectrum below illustrates that **Lithium and Strontium are present** in the mixture as the lines match. There is no **cadmium** present in this mixture.



**Determining concentration of metal ions**

The concentration of metal ions present in a sample can also be determined by the spectrometer. The intensity (or absorbance) of light is measured with a specific wavelength (characteristic of a particular metal ion). The machine can be calibrated using solutions of the metal ion of known concentration and so enabling the unknown concentration to be calculated.

**Use** Monitor water for metals ions (like aluminium, calcium, mercury, lead, cadmium etc).