3.1 Acids

YOUR NOTES

CONTENTS

3.1.1 ACIDS & BASES

Defining Acids & Bases

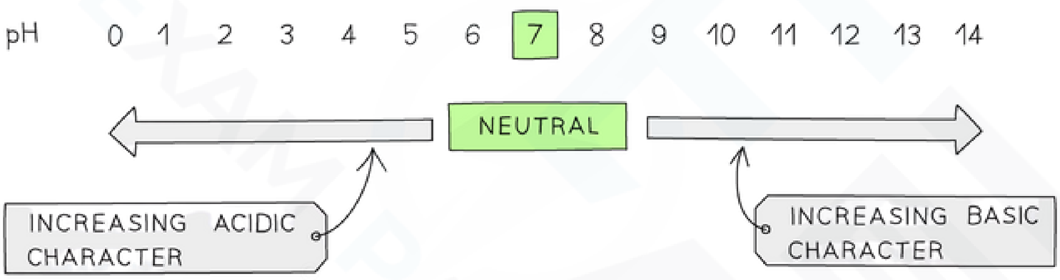
- When **acids** are added to water, they form positively charged **hydrogen ions** (H+)
- The presence of H+ ions is what makes a solution acidic
- When **alkalis** are added to water, they form negative hydroxide ions (OH–)
- The presence of the OH– ions is what makes the aqueous solution an alkali
- The pH scale is a numerical scale which is used to show how **acidic** or **alkaline** a solution is, in other words it is a measure of the amount of the ions present in solution

3.1 Acids

YOUR NOTES

The pH Scale

- The pH scale goes from 0 – 14 (extremely acidic substances can have values of below 0)
- All acids have pH values of **below** 7, all alkalis have pH values of **above** 7
- The **lower** the pH then the **more acidic** the solution is
- The **higher** the pH then the **more alkaline** the solution is
- A solution of pH 7 is described as being **neutral**



The pH scale showing acidity, neutrality and alkalinity

Indicators

- Two colour indicators are used to distinguish between acids and alkalis
- Many plants contain substances that can act as indicators and the most common one is **litmus** which is extracted from lichens
- Synthetic indicators are organic compounds that are sensitive to changes in acidity and appear different colours in acids and alkalis
- **Phenolphthalein** and **methyl orange** are synthetic indicators frequently used in acid-alkali titrations

Two Colour Indicators Table

Indicator	Colour in acid	Colour in alkali
Litmus	red	blue
Phenolphthalein	colourless	pink
Methyl orange	red	yellow

3.1 Acids

YOUR NOTES

- Synthetic indicators are used to show the endpoint in titrations as they have a very sharp change of colour when an acid has been neutralised by an alkali and vice-versa
- Litmus is not suitable for titrations as the colour change is not sharp and it goes through a purple transition colour in neutral solutions making it difficult to determine an endpoint
- Litmus is very useful as an an indicator paper and comes in red and blue versions, for dipping into solutions or testing gases

3.1 Acids

YOUR NOTES

3.1.2 HYDROGEN IONS & PH

Higher Tier Only

Hydrogen Ions & pH

- We have already seen that acids are substances that contain hydrogen ions in solution
- The more hydrogen ions the stronger the acid, but the lower the pH
- The higher the concentration of hydroxide ions in a solution the higher the pH
- So pH is a measure of the concentration of H+ ions in solution, but they have an inverse relationship

3.1 Acids

YOUR NOTES
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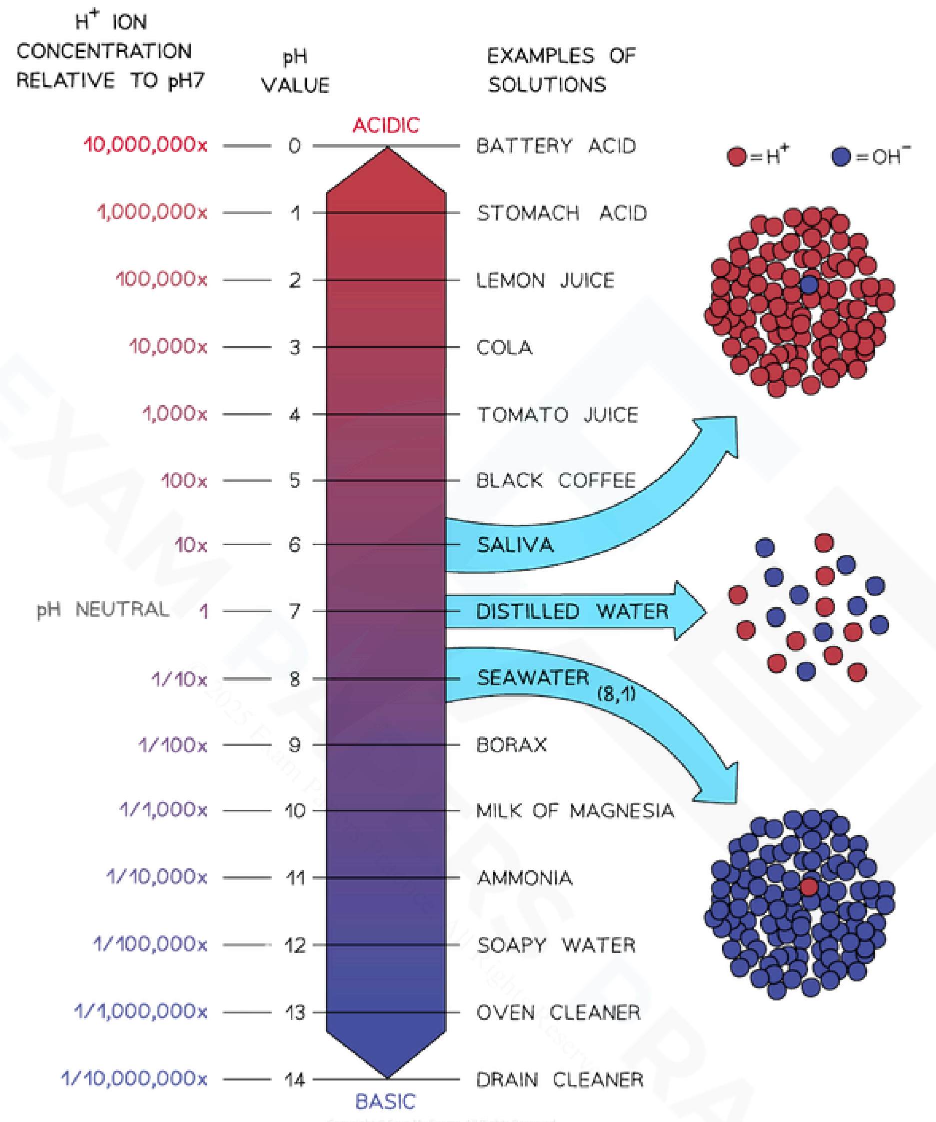


The pH scale is logarithmic so that each change in pH is a tenfold change in hydrogen ion concentration

- The pH scale is **logarithmic**, meaning that each change of 1 on the scale represents a change in concentration by a **factor** of **10**
- Therefore an acid with a pH of 3 has ten times the concentration of H+ ions than an acid of pH 4
- An acid with a pH of 2 has 10 x 10 = 100 times the concentration of H+ ions than an acid with a pH of 4
- From this we can summarize that for two acids of **equal concentration**, where one is **strong** and the other is **weak**, then the strong acid will have a lower pH due to its capacity to dissociate more and hence put more H+ ions into solution than the weak acid

3.1 Acids

YOUR NOTES

Exam Tip

Acid strength is reflected in how many hydrogen ions are in solution. The more hydrogen ions the lower the pH and vice-versa.

3.1 Acids

YOUR NOTES

3.1.3 CORE PRACTICAL: INVESTIGATING PH

Core Practical: Investigating pH

Aim:

- To investigate the changes in pH of a fixed volume of dilute HCl on addition of varying amounts of a solid base

Materials:

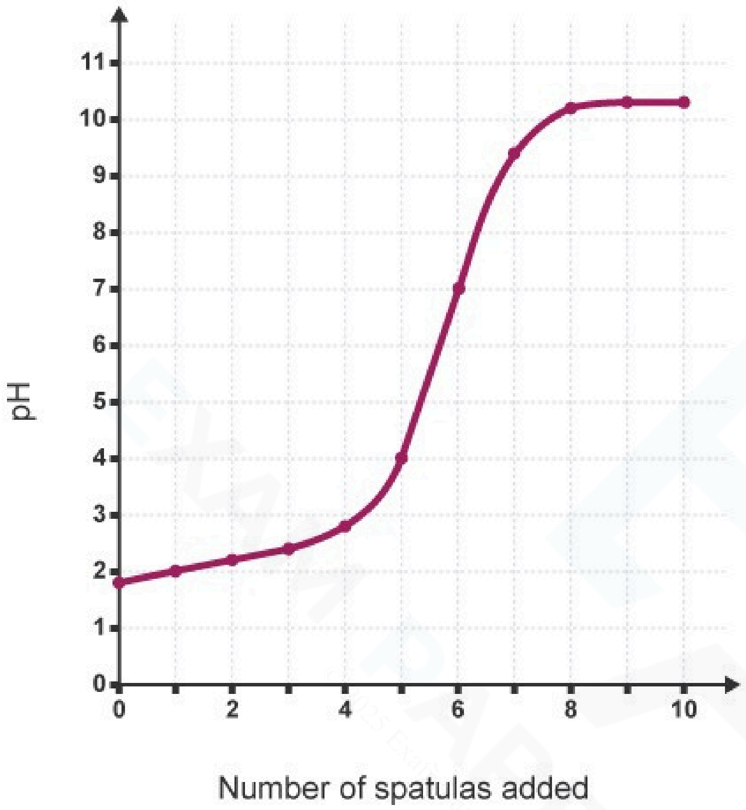
- Dilute HC*l (0.5M or 1M),* solid base such as CaO or Ca(OH)2
- Conical flask, 25 cm3 or 50 cm3 volumetric pipette, glass rod
- Spatula and weighing boat
- pH probe or Universal Indicator paper

Method:

- Use a pipette to measure a fixed volume of dilute HC*l* into a conical flask
- Add one spatula of calcium oxide or calcium hydroxide to the flask and swirl
- When all the base has reacted record the pH of the solution
- If using U.I. paper use the glass rod to extract a sample from the flask
- Repeat for different numbers of spatula (1-10) of solid but the same volume of HC*l*
- Record your results neatly in table format

Analysis of results:

- Plot a graph of the amount of the base on the X-axis against the pH recorded on the Y-axis
- The resulting graph should look something like the one below



Investigating the change in pH during neutralisation of an acid

Conclusion:

- The graph indicates a sudden change in pH which corresponds to the vertical section of the graph
- This indicates that the more solid base is added the higher the pH, therefore the base is neutralising the acid
- From the sample graph it can be seen that 6 spatulas of the base are required to completely neutralise the acid

3.1 Acids

YOUR NOTES

3.1.4 ACID STRENGTH & CONCENTRATION

Higher Tier Only

Acid Strength & Concentration

Strong & Weak

- Acids can be either **strong** or **weak,** depending on how many ions they produce when they dissolve in water
- When added to water, acids ionise or **dissociate** to produce H+ ions:

Hydrochloric acid: $HCl \rightarrow H^+ + Cl^-$

Nitric acid: $HNO_3 \rightarrow H^+ + NO_3^-$

- Strong acids such as HCl and H2SO4 dissociate **completely** in water, producing solutions with a high concentration of H+ ions and thus a very **low pH**
- Weak acids such as ethanoic acid, CH3COOH and hydrofluoric acid, HF only partially ionize in water, producing solutions of pH values between 4 – 6
- This data is summarized in the table below:

Strong & Weak Acids Table

	High Concentration of H^+	Low Concentration of H^+
pH	Low pH	Higher pH (but still under 7)

- For weak acids there is an **equilibrium** set-up between the molecules and their ions once they have been added to water
- Propanoic acid for example dissociates as follows:

$CH_3CH_2COOH \rightleftharpoons H^+ + CH_3CH_2COO^-$

- The $\rightleftharpoons$ symbol indicates that the process is reversible, as the products can react together forming the original reactants
- The equilibrium lies to the **left**, meaning there is a high concentration of intact acid molecules and therefore a low concentration of ions in solution, hence the pH is that of a weak acid and closer to 7 than a strong acid

3.1 Acids

YOUR NOTES

Concentrated & Dilute

- A solution is formed when a solute is dissolved in a solvent
- A **dilute** solution contains a **small** amount of solute in a given volume of solution
- A **concentrated** solution contains a **large** amount of solute in a given volume of solution
- A concentrated solution of either an acid or a base is one that contains a **high number** of acid or base **molecules** per dm3 of solution
- A dilute acid or base solution is therefore one that has much fewer acid or base molecules per dm3 of solution

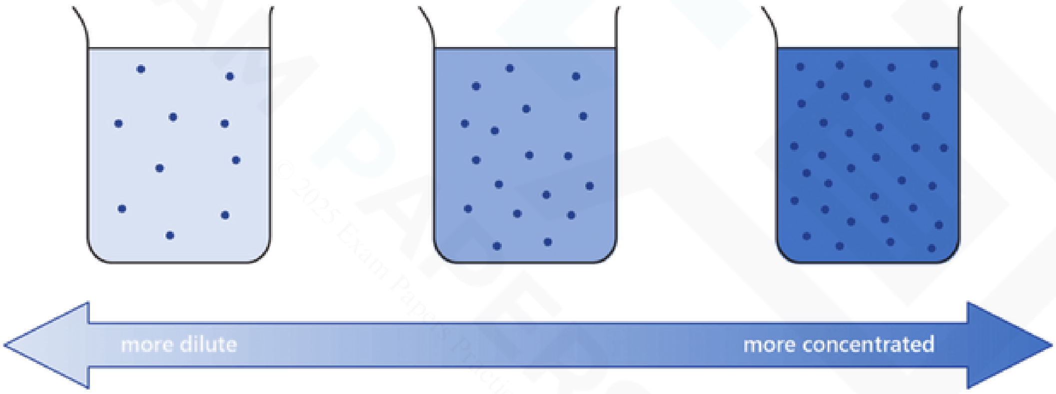


Diagram illustrating how the concentration of a solution increases as more solute is added

Exam Tip

The terms strong and weak refer to the ability to dissociate and not concentration. A dilute solution of a strong acid can have a lower pH than a concentrated solution of a weak acid, due to the stronger acid undergoing complete dissociation.

3.1 Acids
3.1.5 BASES

YOUR NOTES

Bases

What makes a base act like a base?

- Bases are substances which can neutralise an acid, forming a salt and water
- The term **base** and **alkali** are not the same
- A base which is water-soluble is referred to as an **alkali**
  - So, all alkalis are bases, but not all bases are alkalis
- Alkalis have pH values of above 7
- In basic (alkaline) conditions red litmus paper turns **blue**
- Bases are usually **oxides**, **hydroxides** or **carbonates** of metals
- The presence of the OH– ions is what makes the aqueous solution an alkali
- One unusual base is ammonia solution
  - When ammonia reacts with water it produces hydroxide ions

Some Common Alkalis and the Ions they Contain

Name of alkali	Formula	Ions formed in water
Sodium hydroxide	$NaOH$	$Na^+ + OH^-$
Potassium hydroxide	KOH	$K^+ + OH^-$
Aqueous ammonia	NH_3 (+ H_2O)	$NH_4^+ + OH^-$

Exam Tip

Aqueous ammonia and ammonium hydroxide are the same thing. When ammonia gas dissolves in water it forms ammonium hydroxide. Be careful to use the correct terminology: **ammonia** is the gas, **NH_3**, **ammonium** is the ion present in ammonium compounds, **$NH4^+$**

3.1 Acids

YOUR NOTES

3.1.6 REACTIONS OF ACIDS

Reactions of Acids

Reactions of acids with metals

- Only metals **above hydrogen** in the reactivity series will react with dilute acids
- The more reactive the metal then the more vigorous the reaction will be
- Metals that are placed high on the reactivity series such as potassium and sodium are very dangerous and react **explosively** with acids
- When acids react with metals they form a **salt** and **hydrogen gas:**
- The general equation is:

metal + acid → salt + hydrogen

- Some examples of metal-acid reactions and their equations are given below:

Acid-Metals Reactions Table

Acid	Sulfuric Acid	Hydrochloric Acid
Magnesium	$Mg + H_2SO_4 \longrightarrow MgSO_4 + H_2$	$Mg + 2HCl \longrightarrow MgCl_2 + H_2$
Zinc	$Zn + H_2SO_4 \longrightarrow ZnSO_4 + H_2$	$Zn + 2HCl \longrightarrow ZnCl_2 + H_2$
Iron	$Fe + H_2SO_4 \longrightarrow FeSO_4 + H_2$	$Fe + 2HCl \longrightarrow FeCl_2 + H_2$

- In general, we can summarise the reaction of a metal that forms a +2 ion as follows:

Acids-Metals Summary Table

Acid	Name of products	Equation for reaction
Hydrochloric acid	Metal chloride and hydrogen	$M + 2HCl \longrightarrow MCl_2 + H_2$
Sulfuric acid	Metal sulfate and hydrogen	$M + H_2SO_4 \longrightarrow MSO_4 + H_2$

3.1 Acids

YOUR NOTES

Reaction of acids with oxides & hydroxides

- When an acid reacts with an oxide or hydroxide, a **neutralisation** reaction occurs
- Metal **oxides** and metal **hydroxides** act as **bases**
- In all acid-base neutralisation reactions, a **salt** and **water** are produced:

acid + base → salt + water

- The identity of the salt produced depends on the **acid** used and the **positive ions** in the base
- Hydrochloric acid produces **chlorides**, sulfuric acid produces **sulfate** salts and nitric acid produces **nitrates**
- The following are some specific examples of reactions between acids and metal oxides / hydroxides:

2HCl + CuO → CuCl2 + H2O

H2SO4 + 2NaOH → Na2SO4 + 2H2O

HNO3 + KOH → KNO3 + H2O

- In general, we can summarise the reaction of metals and bases as follows:

Acids and Metals Oxides or Hydroxides Summary Table

Acid	Name of Products	Equation for Reaction
Hydrochloric Acid	Metal Chloride and Water	$MOH + HCl \longrightarrow MCl + H_2O$
Sulfuric Acid	Metal Sulfate and Water	$MO + H_2SO_4 \longrightarrow MSO_4 + H_2O_4$
Nitric Acid	Metal Nitrate and Water	$MO + HNO_3 \longrightarrow MNO_3 + H_2O$

3.1 Acids

YOUR NOTES

Reactions of Acids with Metal Carbonates

- Acids will react with metal carbonates to form the corresponding metal **salt**, **carbon dioxide** and **water**
- These reactions are easily distinguishable from acid – metal oxide/hydroxide reactions due to the presence of **effervescence** caused by the carbon dioxide gas

Acids & Metal Carbonates Reactions Table

Acid	Name of Products	Equation for Reaction
Hydrochloric Acid	Metal Chloride, Carbon Dioxide and Water	$MCO_3 + 2HCl \longrightarrow MCl_2 + CO_2 + H_2O$
Sulfuric Acid	Metal Sulfate, Carbon Dioxide and Water	$MCO_3 + H_2SO_4 \longrightarrow MSO_4 + CO_2 + H_2O$
Nitric Acid	Metal Nitrate, Carbon Dioxide and Water	$MCO_3 + 2HNO_3 \longrightarrow M(NO_3)_2 + CO_2 + H_2O$

- The following are some specific examples of reactions between acids and metal carbonates:

2HCl + Na2CO3 → 2NaCl + H2O + CO2

H2SO4 + CaCO3→ CaSO4 + H2O + CO2

Exam Tip

If in an acid-base reaction there is effervescence produced then the base must be a metal carbonate which produces carbon dioxide gas.

3.1 Acids

YOUR NOTES

Neutralisation

- The chemistry of neutralisation reactions can be explained using ionic equations
- Ionic equations are used to show only the particles that chemically participate in a reaction
- The other ions present are not involved and are called **spectator ions**
- For example the neutralisation reaction between hydrochloric acid and sodium hydroxide:

$$HCl + NaOH \rightarrow NaCl + H_2O$$

- If we write out all of the **ions** present in the equation we get:

$$H^+ + Cl^- + Na^+ + OH^- \rightarrow Na^+ + Cl^- + H_2O$$

- The **spectator** ions are thus Na^+ and Cl^-.
- Removing these from the previous equation leaves the overall net ionic equation:

$$H^+ + OH^- \rightarrow H_2O$$

- The H^+ ions come from the acid and the OH^- ions come from the base, both combine to form the product water molecules
- This ionic equation is the same for **all acid-base neutralisation**

Exam Tip

Remember that although acids react with metals to form salts, that reaction is not neutralisation, but it counts as a redox reaction.

YOUR NOTES

3.1.7 TEST FOR HYDROGEN & CARBON DIOXIDE

Test for Hydrogen & Carbon Dioxide

Testing Hydrogen

- The test for hydrogen consists of holding a **burning splint** held at the open end of a test tube of gas
- If the gas is hydrogen it burns with a loud "**squeaky pop**" which is the result of the rapid combustion of hydrogen with oxygen to produce water
- Be sure not to insert the splint right into the tube, just at the mouth, as the gas needs air to burn

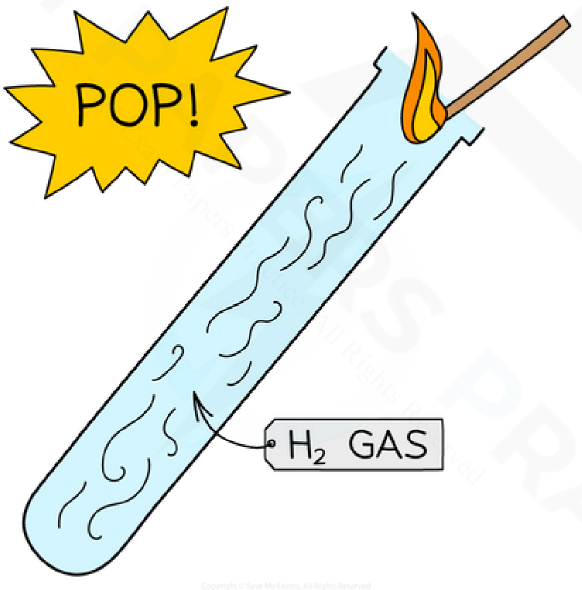


Diagram showing the test for hydrogen gas

Exam Tip

It is easy to confuse the tests for hydrogen and oxygen. Try to remember that a lig**H**ted splint has a **H** for **H**ydrogen, while a gl**O**wing splint has an **O** for **O**xygen.

3.1 Acids

Testing Carbon Dioxide

- The test for carbon dioxide involves bubbling the gas through an aqueous solution of **calcium hydroxide** (limewater)
- If the gas is carbon dioxide, the limewater turns **milky** or **cloudy**

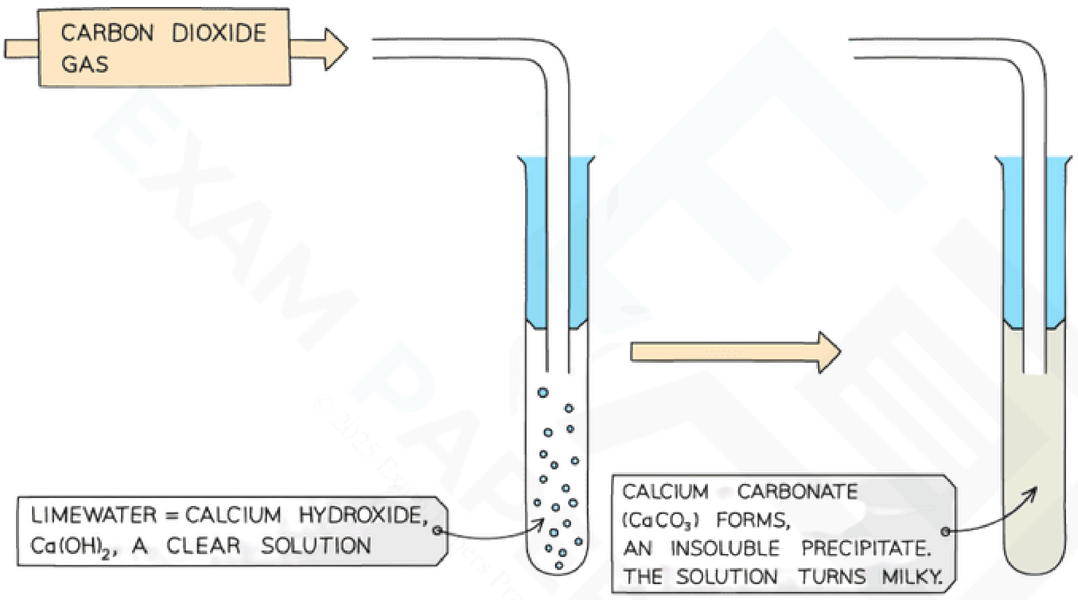


Diagram showing the test for carbon dioxide gas

Exam Tip

Sometimes students think that extinguishing a burning splint indicates carbon dioxide gas. However, while it is a property of carbon dioxide, other gases, such as nitrogen, will also do this, so the test is not definitive and should not be quoted in an exam answer.

3.1 Acids

YOUR NOTES

3.1.8 CORE PRACTICAL: PREPARING COPPER SULFATE

Core Practical: Preparing Copper Sulfate

Aim:

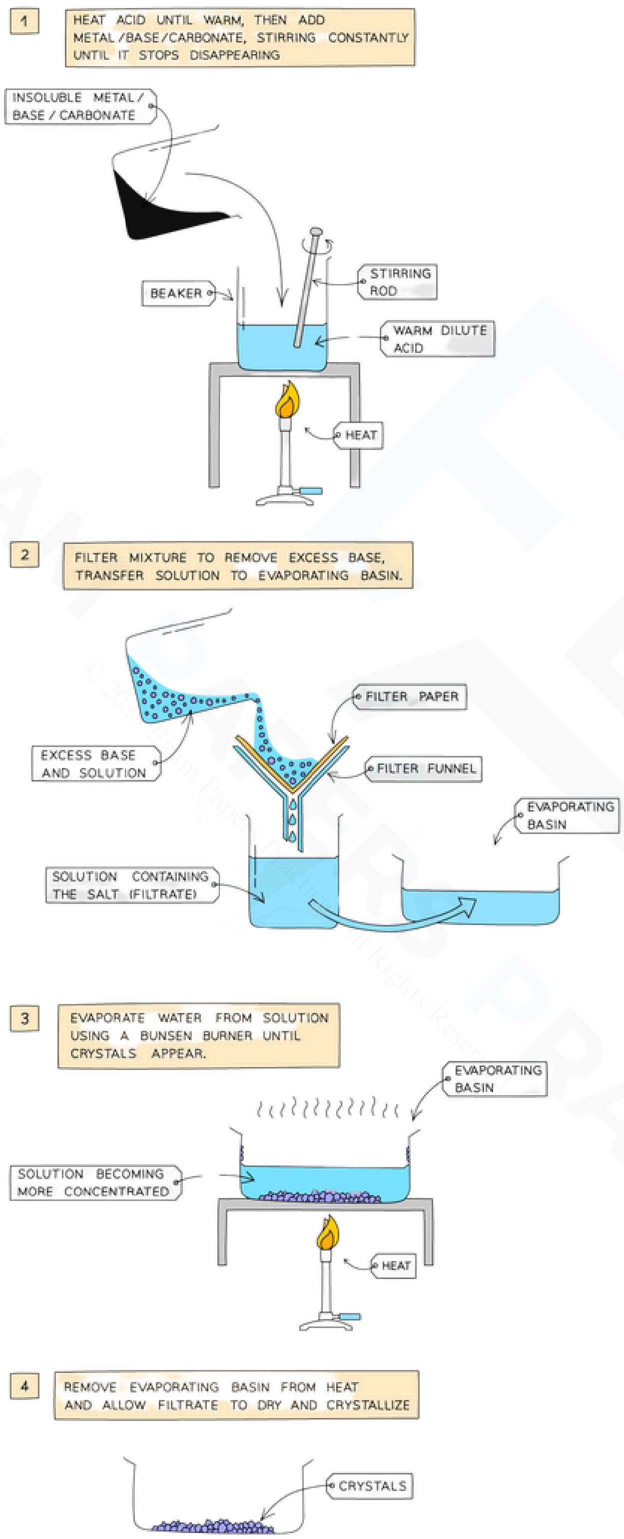
To prepare a pure, dry sample of hydrated copper(II) sulfate crystals

Materials:

- 1.0 mol / dm3 dilute sulfuric acid
- Copper(II) oxide
- Spatula & glass rod
- Measuring cylinder & 100 cm3 beaker
- Bunsen burner
- Tripod, gauze & heatproof mat
- Filter funnel & paper, conical flask
- Evaporating basin and dish.

3.1 Acids

YOUR NOTES



The preparation of copper(II) sulfate by the insoluble base method

3.1 Acids

YOUR NOTES

Practical Tip:
The base is added in excess to use up all of the acid, which would become dangerously concentrated during the evaporation and crystallisation stages

Method:

1. Add 50 cm3 dilute acid into a beaker and warm gently using a Bunsen burner
2. Add the copper(II) oxide slowly to the hot dilute acid and stir until the base is in excess (i.e. until the base stops dissolving and a suspension of the base forms in the acid)
3. Filter the mixture into an evaporating basin to remove the excess base
4. Gently heat the solution in a water bath or with an electric heater to evaporate the water and to make the solution saturated
5. Check the solution is saturated by dipping a cold glass rod into the solution and seeing if crystals form on the end
6. Leave the filtrate in a warm place to dry and crystallise
7. Decant excess solution and allow the crystals to dry

Results:
Hydrated copper(II) sulfate crystals should be bright blue and regularly shaped

Exam Tip

Make sure you learn the names of all the laboratory apparatus used in the preparation of salts.

3.1 Acids

YOUR NOTES

3.1.9 PREPARE A SALT BY TITRATION

Prepare a Salt by Titration

- If salts are prepared from an acid and a soluble reactant then a **titration technique** must be used
- In a titration, the exact volume of acid and soluble reactant are mixed in the correct proportions so that all that remains is the salt and water

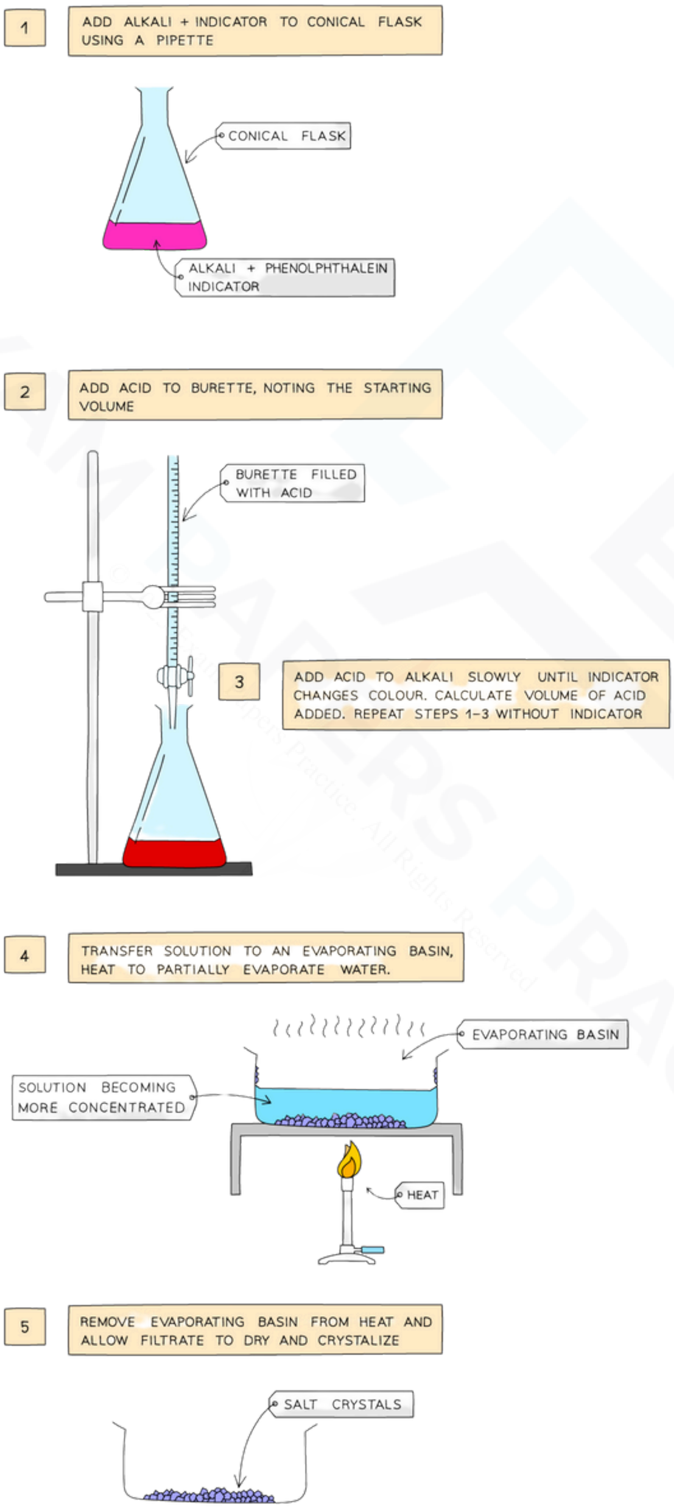
Preparing a Salt by Titration

Aim:

To prepare a sample of a dry salt starting from an acid and an alkali

3.1 Acids

Diagram:



3.1 Acids

Diagram showing the apparatus needed to prepare a salt by titration

Method:

- Use a pipette to measure the alkali into a conical flask and add a few drops of indicator (phenolphthalein or methyl orange)
- Add the acid into the burette and note the starting volume
- Add the acid very slowly from the burette to the conical flask until the indicator changes to appropriate colour
- Note and record the final volume of acid in burette and calculate the volume of acid added (starting volume of acid – final volume of acid)
- Add this same volume of acid into the same volume of alkali without the indicator
- Heat to partially evaporate, leaving a saturated solution
- Leave to crystallise decant excess solution and allow crystals to dry

Results:

A dry sample of a salt is obtained

Exam Tip

When evaporating the solution some water is left behind to allow for water of crystallisation in some salts and also to prevent the salt from overheating and decomposing.

3.1 Acids

YOUR NOTES

3.1.10 SOLUBILITY RULES

Solubility Rules

- Ionic compounds are generally soluble in water compared to covalent substances, but there are exceptions
- A knowledge of the solubility of ionic compounds helps us to determine the most appropriate method for the preparation of salts
- The solubility of common ionic compounds is shown below:

Solubility of Ionic Compounds Table

Soluble	Insoluble
Compounds of sodium, potassium, and ammonium	–
All nitrates	–
All chlorides, except...	...silver and lead (II)
All sulfates except...	...barium, calcium, and lead (II)
Sodium, potassium, andd ammonium carbonates	All other carbonates
Sodium, potassium, and calcium	All other hydroxides

- Calcium hydroxide is slightly soluble in water

Exam Tip

Calcium hydroxide solution is more commonly know as limewater and is used to test for carbon dioxide.

3.1 Acids

YOUR NOTES

Predicting Precipitates

- Some salts can be extracted by mining but others need to be prepared in the laboratory
- How the salt is made in the laboratory depends on whether the salt being formed is **soluble** or **insoluble** in water
- To do this the balanced equation is written down to determine the identify of the salt product
- Then check the solubility of the salt using the **solubility** table
- If it is **soluble** in water, then it can be prepared by **titration**
- If it is **insoluble** then it can be prepared by **precipitation**
- For example a silver nitrate solution is mixed with a sodium chloride solution:

AgNO3 (aq) + NaC*l* (aq) → AgC*l* (s) + NaNO3 (aq)

- From the table both AgNO3 and NaC*l* are water soluble but AgC*l*, silver chloride, is not and hence forms a **precipitate**

Exam Tip

The precipitation reaction by combining two soluble salts is also known as a double decomposition or double displacement reaction.

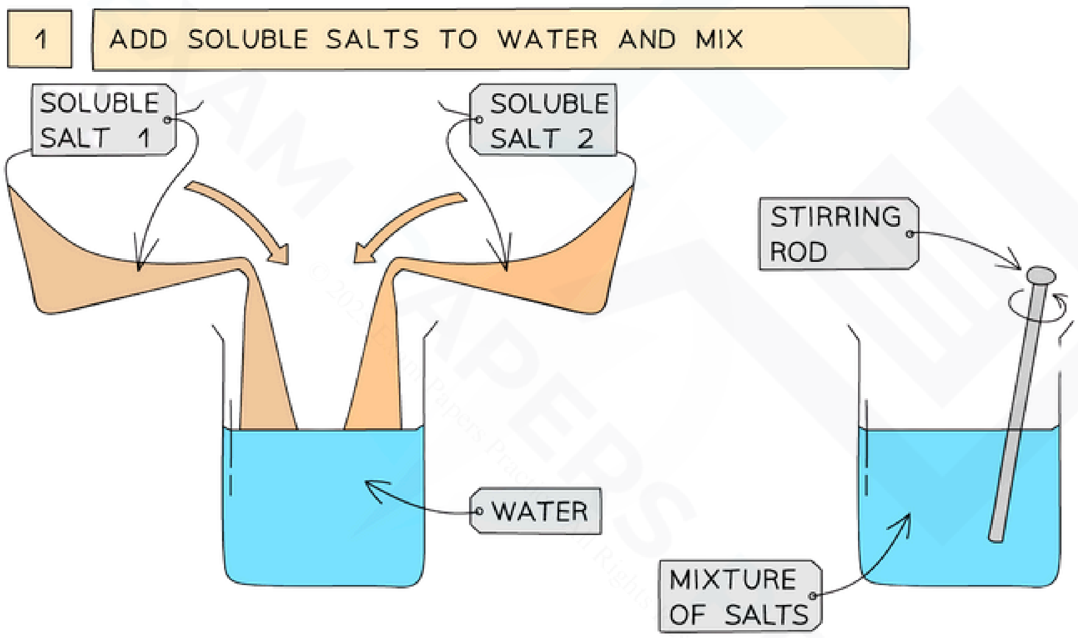
3.1 Acids

3.1.11 PREPARING AN INSOLUBLE SALT

Preparing an Insoluble Salt

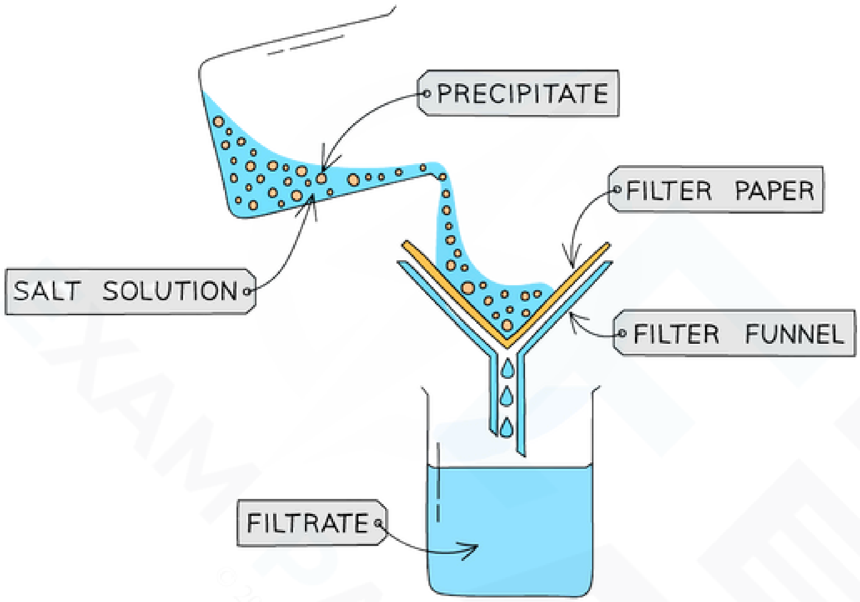
Aim:
To prepare a dry sample of an insoluble salt, lead(II) sulfate

Diagram:

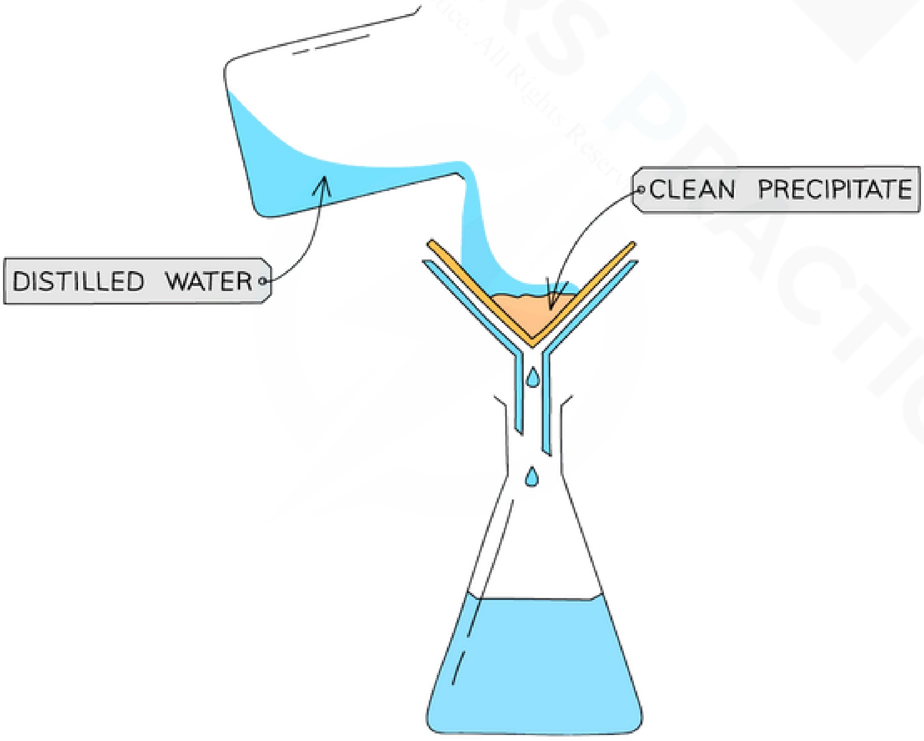


3.1 Acids

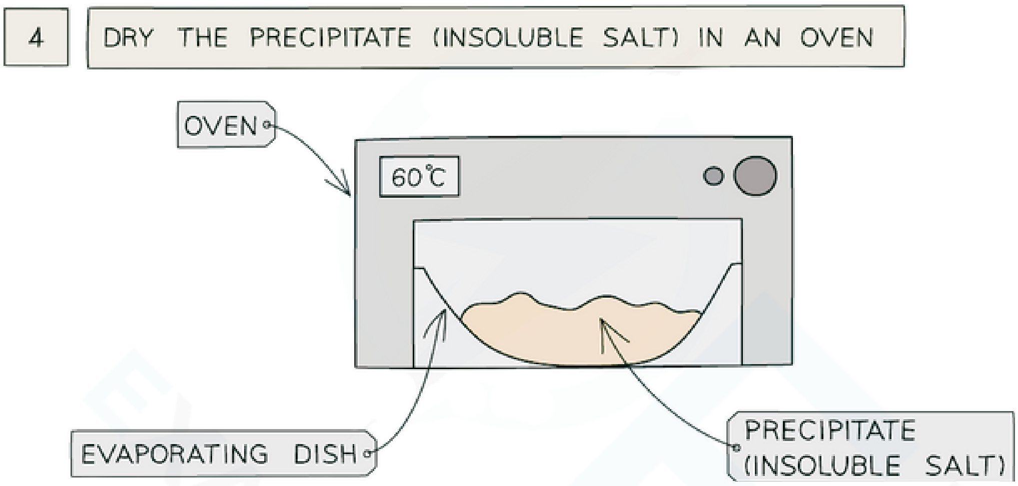
2 FILTER TO REMOVE PRECIPITATE FROM THE MIXTURE



3 WASH THE PRECIPITATE WITH DISTILLED WATER TO REMOVE TRACES OF SOLUTION



3.1 Acids



The preparation of lead(II)sulfate by precipitation from two soluble salts

Method:

- Measure out 25 cm3 of 0.5 mol dm3 lead(II) nitrate solution and add it to a small beaker
- Measure out 25 cm3 of 0.5 mol dm3 of potassium sulfate add it to the beaker and mix together using a stirring rod
- Filter to remove the precipitate from mixture
- Wash the filtrate with distilled water to remove traces of other solutions
- Leave in an oven to dry

Soluble salt 1 = lead(II) nitrate

Soluble salt 2 = potassium sulfate

Equation for the reaction:

$$Pb(NO_3)_2\ (aq) + K_2SO_4\ (aq) \rightarrow PbSO_4\ (s) + 2KNO_3\ (aq)$$

lead(II) nitrate + potassium sulfate → lead(II) sulfate + potassium nitrate

Exam Tip

Care should be taken with handling lead salts as they are toxic.