IGCSE Chemistry CIE

YOUR NOTES

7. Acids, Bases & Salts

CONTENTS

© 2015-2021 Save My Exams, Ltd. · Revision Notes, Topic Questions, Past Papers

YOUR NOTES
⬇

7.1 The Characteristic Properties of Acids & Bases

7.1.1 Properties of Acids & Bases

Properties of Acids

- Acids have pH values of below 7, have a sour taste (when edible) and are corrosive
- Acids are substances that can neutralise a base, forming a salt and water
- When **acids** are added to water, they form positively charged **hydrogen ions** (H^+)
- The presence of H^+ ions is what makes a solution acidic

Example: Hydrochloric Acid

$$HCl\ (aq) \rightarrow H^+\ (aq) + Cl^-\ (aq)$$

Typical reactions of acids

Acids and metals

- Only metals **above hydrogen** in the reactivity series will react with dilute acids.
- When acids react with metals they form a salt and hydrogen gas:

Acid + Metal → Salt + Hydrogen

- The name of the salt is related to the name of the acid used, as it depends on the **anion** within the acid.
- Examples of the names of salts from specific acids and metals are:

Acid	Name of products	Equation for reaction
Hydrochloric acid	Magnesium chloride and hydrogen	$Mg + 2HCl \longrightarrow MgCl_2 + H_2$
Sulfuric acid	Magnesium sulfate and hydrogen	$Mg + H_2SO_4 \longrightarrow MgSO_4 + H_2$
Nitric acid	Magnesium nitrate and hydrogen	$Mg + 2HNO_3 \longrightarrow Mg(NO_3)_2 + H_2$

Copyright © Save My Exams. All Rights Reserved

Acids with bases

- Metal **oxides** and metal **hydroxides** (**alkalis**) can act as **bases**
- When they react with acid, a **neutralisation** reaction occurs
- In all acid-base neutralisation reactions, **salt** and **water** are produced

Acid + Base → Salt + Water

- Examples of reactions between acids and bases:

Acid	Name of products	Equation for reaction
Hydrochloric acid	Magnesium chloride and water	$Mg(OH)_2 + 2HCl \longrightarrow MgCl_2 + 2H_2O$
Sulfuric acid	Magnesium sulfate and water	$MgO + H_2SO_4 \longrightarrow MgSO_4 + H_2O$
Nitric acid	Magnesium nitrate and water	$Mg(OH)_2 + 2HNO_3 \longrightarrow Mg(NO_3)_2 + H_2O$

Copyright © Save My Exams. All Rights Reserved

Acids with metal carbonates

© 2015-2021 Save My Exams, Ltd. · Revision Notes, Topic Questions, Past Papers

YOUR NOTES
⬇

- Acids will react with metal carbonates to form the corresponding metal salt, carbon dioxide and water:

Acid + Metal Carbonate → Salt + Carbon Dioxide + Water

- Examples of reactions between acids and carbonates:

Acid	Name of products	Equation for reaction
Hydrochloric acid	Magnesium chloride, carbon dioxide and water	$MgCO_3 + 2HCl \longrightarrow MgCl_2 + CO_2 + H_2O$
Sulfuric acid	Magnesium sulfate, carbon dioxide and water	$MgCO_3 + H_2SO_4 \longrightarrow MgSO_4 + CO_2 + H_2O$
Nitric acid	Magnesium nitrate, carbon dioxide and water	$MgCO_3 + 2HNO_3 \longrightarrow Mg(NO_3)_2 + CO_2 + H_2O$

Copyright © Save My Exams. All Rights Reserved

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
- **Thymolphthalein** 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
Thymolphthalein	colourless	blue
Methyl orange	red	yellow

Copyright © Save My Exams. All Rights Reserved

- 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 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 indicator paper and comes in red and blue versions, for dipping into solutions or testing gases

© 2015-2021 Save My Exams, Ltd. · Revision Notes, Topic Questions, Past Papers

YOUR NOTES

Properties of Bases & Alkalis

- Bases have pH values of above 7
- A base which is water-soluble is referred to as an **alkali**
- In basic (alkaline) conditions red litmus paper turns **blue**, methyl orange indicator turns **yellow** and thymolphthalein indicator turns **blue**
- Bases are substances which can neutralise an acid, forming a salt and water
- Bases are usually **oxides** or **hydroxides** of metals
- 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

Example: Sodium Hydroxide

$$NaOH\ (s) \rightarrow Na^+\ (aq) + OH^-\ (aq)$$

Typical reactions of bases

Bases and acids

- When bases react with an acid, a neutralisation reaction occurs
- Acids and bases react together in a neutralisation reaction and produce a salt and water:

Acid + Base → Salt + Water

Examples of reaction between bases and acids:

Acid	Name of products	Equation for reaction
Hydrochloric acid	Magnesium chloride and water	$Mg(OH)_2 + 2HCl \longrightarrow MgCl_2 + 2H_2O$
Sulfuric acid	Magnesium sulfate and water	$MgO + H_2SO_4 \longrightarrow MgSO_4 + H_2O$
Nitric acid	Magnesium nitrate and water	$Mg(OH)_2 + 2HNO_3 \longrightarrow Mg(NO_3)_2 + H_2O$

Copyright © Save My Exams. All Rights Reserved

Alkalis and ammonium salts

- Ammonium salts undergo decomposition when warmed with an alkali
- Even though ammonia is itself a weak base, it is very **volatile** and can easily be displaced from the salt by another alkali
- A salt, water and ammonia are produced

Example:

$$NH_4Cl + NaOH \rightarrow NaCl + H_2O + NH_3$$

- This reaction is used as a chemical test to confirm the presence of the ammonium ion (NH_4^+)
- Alkali is added to the substance with gentle warming followed by the test for ammonia gas using damp red litmus paper
- The damp litmus paper will turn from **red** to **blue** if ammonia is present

© 2015-2021 Save My Exams, Ltd. · Revision Notes, Topic Questions, Past Papers

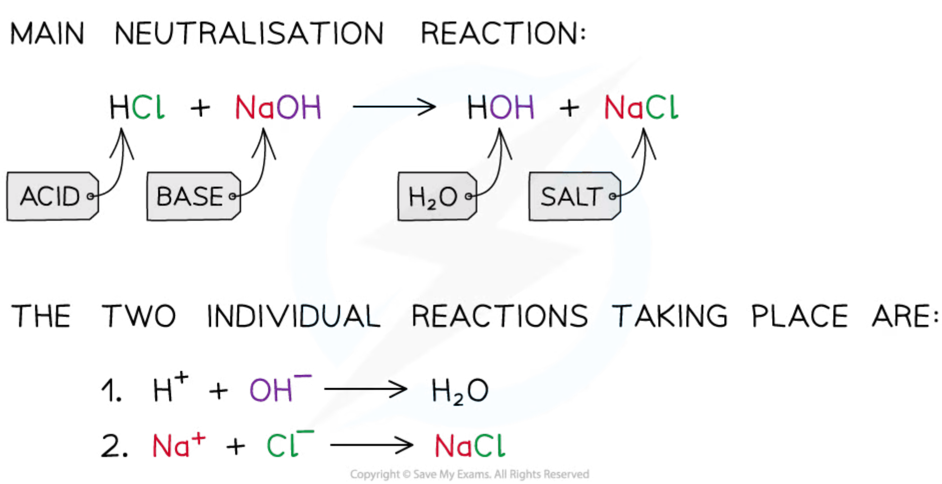
7.1.2 The Ions in Acids & Alkalis

YOUR NOTES

Neutralisation Reactions

- 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 ions present in the solution

- A neutralisation reaction occurs when an acid reacts with an alkali
- When these substances react together in a neutralisation reaction, the **H^+** ions react with the **OH^-** ions to produce water
- For example, when hydrochloric acid is neutralised a sodium chloride and water are produced:



- The net ionic equation of **acid-alkali neutralisations**, and what leads to a neutral solution, since water has a pH of 7, is:

$$\mathbf{H^+ + OH^- \rightarrow H_2O}$$

Exam Tip

Not all reactions of acids are neutralisations. For example, when a metal reacts with an acid, although a salt is produced there is no water formed so it does not fit the definition of neutralisation.

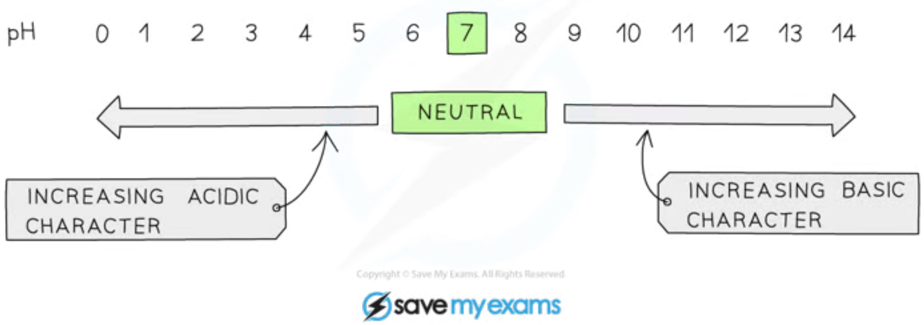
© 2015-2021 Save My Exams, Ltd. · Revision Notes, Topic Questions, Past Papers

Hydrogen Ion Concentration & pH

YOUR NOTES

The pH scale

- The pH scale is a numerical scale which is used to show how **acidic** or **alkaline** a solution is
- It goes from 1 – 14 (extremely acidic substances can have values of below 1)
- 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 with a pH of 7, such as water, is described as being **neutral**

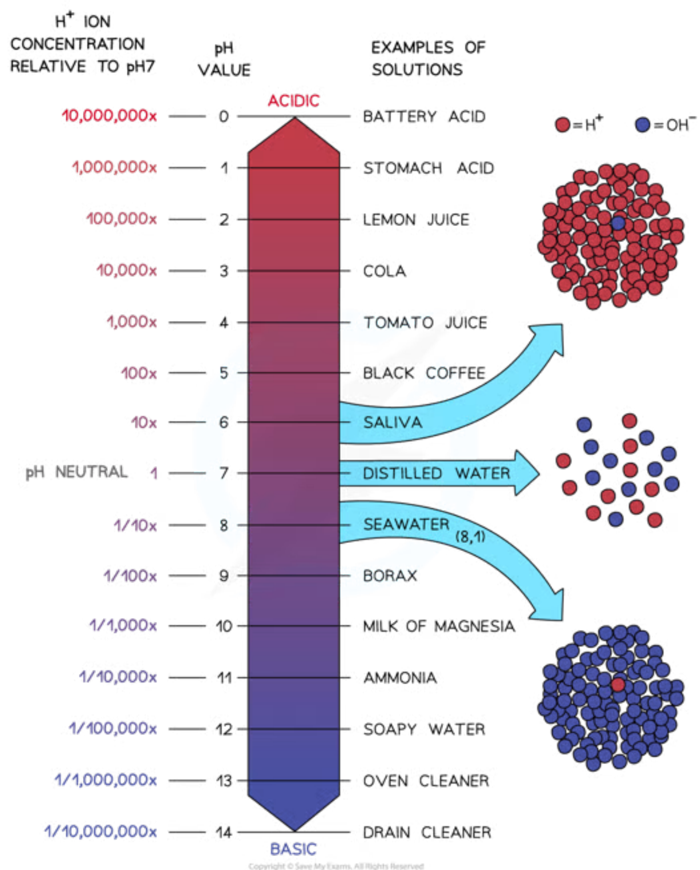


The pH scale showing acidity, neutrality and alkalinity

- 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

© 2015-2021 Save My Exams, Ltd. · Revision Notes, Topic Questions, Past Papers

YOUR NOTES
⬇



- 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

Exam Tip

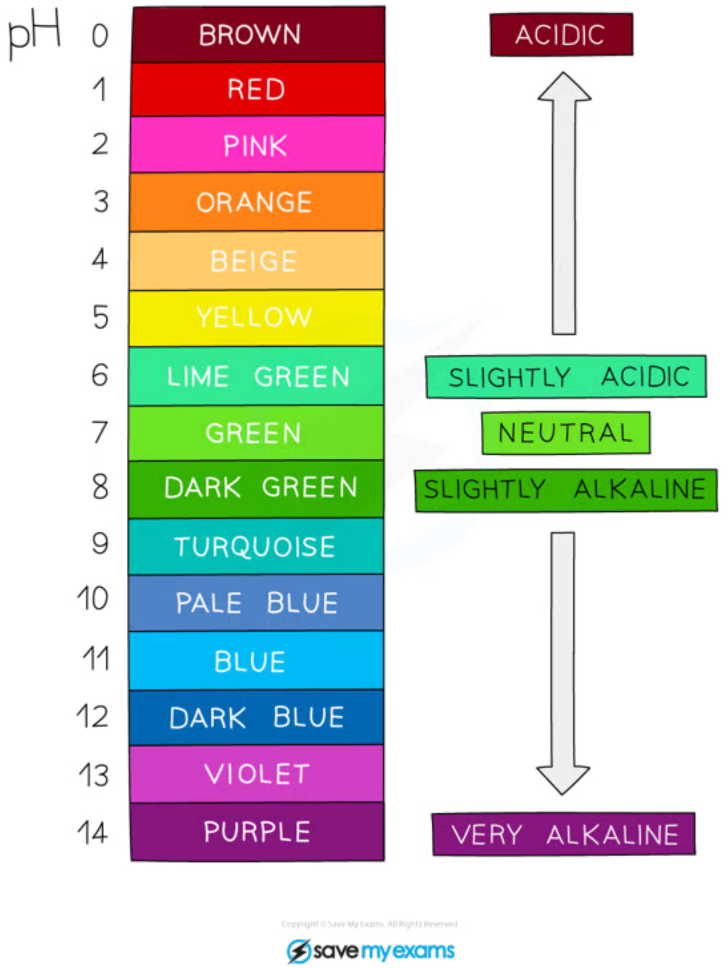
Acid strength is reflected in how many hydrogen ions are in solution. The more hydrogen ions the lower the pH and vice-versa.

Universal indicator

- Universal indicator is a mixture of different **indicators** which is used to measure the pH
- A drop is added to the solution and the colour is matched with a colour chart which indicates the pH which matches specific colours

© 2015-2021 Save My Exams, Ltd. · Revision Notes, Topic Questions, Past Papers

YOUR NOTES
↓



The pH scale with the Universal Indicator colours which can be used to determine the pH of a solution

© 2015-2021 Save My Exams, Ltd. · Revision Notes, Topic Questions, Past Papers

YOUR NOTES

Proton Transfer, Strong & Weak Acids

EXTENDED

Proton transfer

- The earlier definition of an acid and a base can be extended
- In terms of proton transfer, we can further define each substance in how they interact with protons

Acids

- Acids are **proton donors** as they ionise in solution producing protons, which are H^+ ions
- These H^+ ions make the aqueous solution acidic

Bases

- Bases are **proton acceptors** as they accept the protons which are donated by the acid

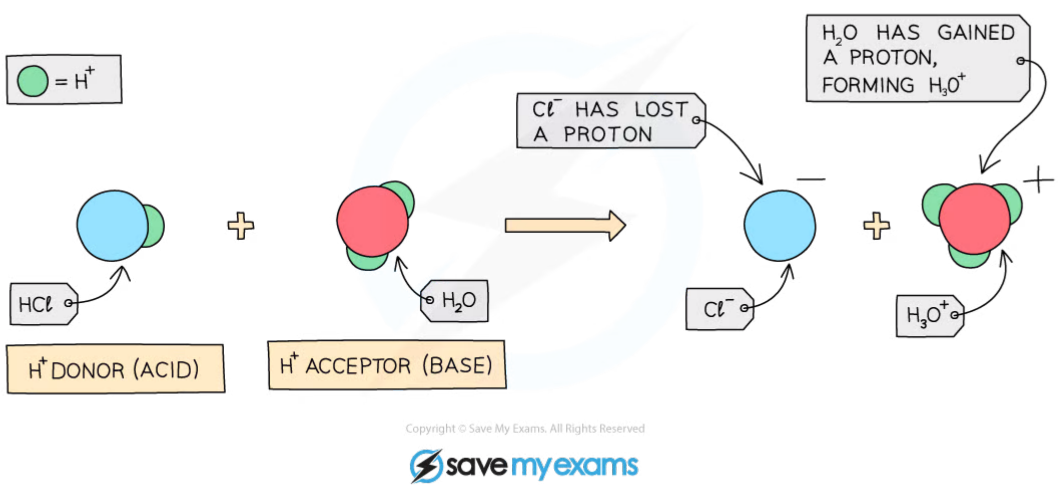


Diagram showing the role of acids and bases in the transfer of protons – here water acts as a base as it accepts a proton

Strong acids

- Acids can be either **strong** or **weak**, depending on how many H^+ ions they produce when dissolved in water
- Strong acids **completely dissociate** (or ionise) in water, producing solutions of a **very low pH**
- Strong acids include HCl and H_2SO_4
- Example of a strong acid: hydrochloric acid

$$HCl\ (aq) \rightarrow H^+\ (aq) + Cl^-\ (aq)$$

© 2015-2021 Save My Exams, Ltd. · Revision Notes, Topic Questions, Past Papers

Weak acids

- Weak acids **partially dissociate** (or ionise) in water and produce pH values which are closer to the **middle** of the pH scale, whilst still being below 7
- Weak acids include organic acids such as ethanoic acid, CH_3COOH
- For weak acids, there is usually an equilibrium set-up between the molecules and their ions once they have been added to water
- Example of a weak acid: propanoic acid

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

- The equilibrium lies to the **left**, indicating a high concentration of intact acid molecules, with a low concentration of H^+ ions in the solution

Effect of concentration on strong and weak acids

- A concentrated solution of an acid is one that contains a **higher number** of acid **molecules** per dm^3 of solution
- It does not necessarily mean that the acid is strong though, as it may be made from a weak acid which does not dissociate completely
- For example a dilute solution of HCl will be more acidic than a concentrated solution of ethanoic acid, since most of the HCl molecules dissociate but very few of the CH_3COOH do

YOUR NOTES
⬇

© 2015-2021 Save My Exams, Ltd. · Revision Notes, Topic Questions, Past Papers

7.1.4 Classifying Oxides

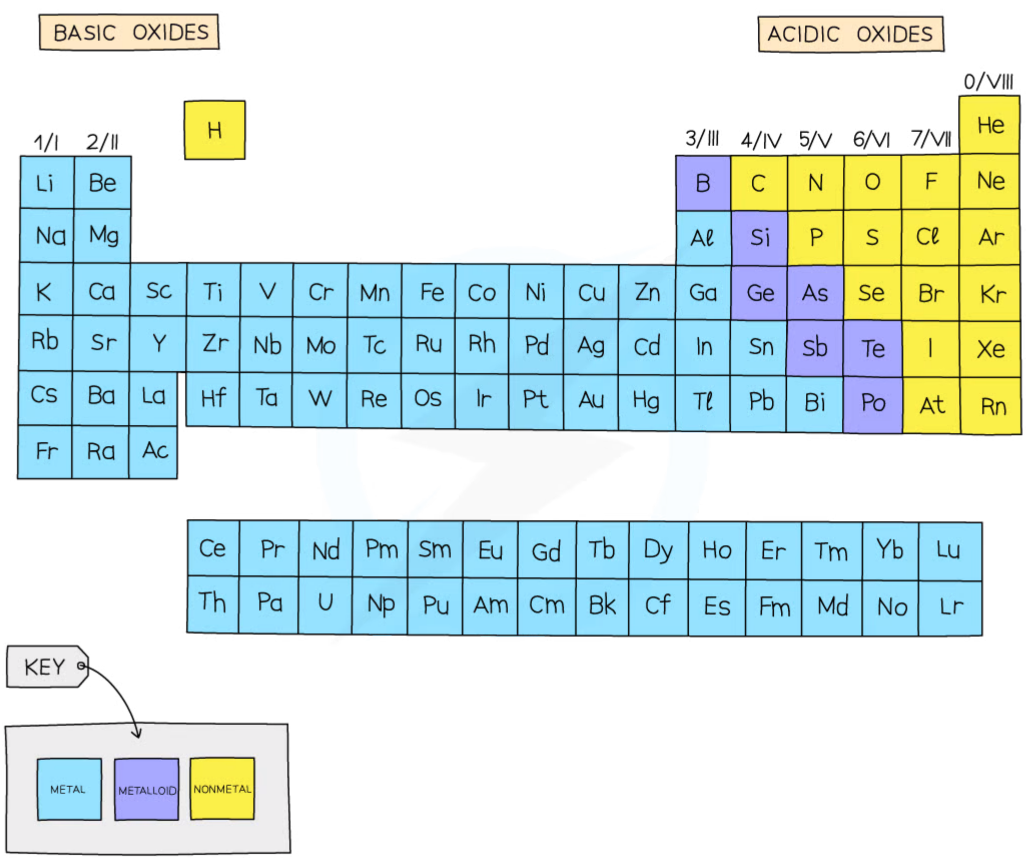
YOUR NOTES

Classifying Oxides

- Oxides are compounds made from one or more atoms of oxygen combined with one other element
- Examples of oxides include: MgO, ZnO, K_2O, CO_2, SO_2, H_2O
- Oxides can be classified based on their acid-base characteristics

Acid and basic oxides

- Acidic and basic oxides have different properties and values of pH
- The difference in their pH stems from whether they are bonded to a **metal** or a **non-metal** element
- The metallic character of the element influences the acidic or basic behaviour of the molecule



Copyright © Save My Exams. All Rights Reserved

Metals form basic oxides while non-metals form acidic oxides

Acidic oxides

- Acidic oxides are formed when a **non-metal** element combines with oxygen

YOUR NOTES
⬇

- They react with bases to form a salt and water
- When dissolved in water they produce an **acidic** solution with a **low** pH
- Common examples include CO_2, SO_2, NO_2 and SiO_2

Basic oxides

- Basic oxides are formed when a **metal** element combines with oxygen
- They react with acids to form a salt and water
- When dissolved in water they produce a **basic** solution with a **high** pH
- Common examples include CuO and CaO

Amphoteric Oxides

EXTENDED

Neutral oxides

- Some oxides do not react with either acids or bases and thus are said to be **neutral**
- Examples include N_2O, NO and CO

Amphoteric oxides

- Amphoteric oxides are a curious group of oxides that can behave as **both acidic** and **basic**, depending on whether the other reactant is an acid or a base
- In both cases salt and water are formed
- Two of the most common amphoteric oxides are **zinc oxide**, **ZnO** and **aluminum oxide**, **Al_2O_3**
- The **hydroxides** of both of these elements also behave amphoterically
- Example of aluminium oxide behaving as a base:

$$Al_2O_3 + 6HCl \rightarrow 2AlCl_3 + 3H_2O$$

- Example of aluminium oxide behaving as an acid:

$$Al_2O_3 + 2NaOH \rightarrow 2NaAlO_2 + H_2O$$

- This acidic and basic behaviour is not easily explained by donating or accepting protons. A separate theory called the Lewis acid-base theory can identify acids or bases in these situations, but is not required for this course

© 2015-2021 Save My Exams, Ltd. · Revision Notes, Topic Questions, Past Papers

YOUR NOTES
⬇

7.2 Preparation of Salts

7.2.1 Preparing Soluble Salts

Preparing Soluble Salts

Salts

- A salt is a compound that is formed when the **hydrogen atom** in an acid is replaced by a **metal**
- For example if we replace the H in HCl with a potassium atom, then the salt potassium chloride is formed, KCl
- Salts are an important branch of chemistry due to the varied and important uses of this class of compounds
- These uses include fertilisers, batteries, cleaning products, healthcare products and fungicides

Naming salts

- The name of salt has two parts
- The first part comes from the **metal**, **metal** oxide or **metal** carbonate used in the reaction
- The second part comes from the **acid**
- The name of the salt can be determined by looking at the reactants
- For example hydrochloric acid always produces salts that end in chloride and contain the **chloride** ion, Cl^-
- Other examples:
  - **Sodium** hydroxide reacts with hydro**chloric** acid to produce sodium chloride
  - **Zinc** oxide reacts with **sulf**uric acid to produce **zinc sulfate**

Preparing salts

- Some salts can be extracted by mining but others need to be prepared in the laboratory
- The method used depends on the solubility of the salt being prepared

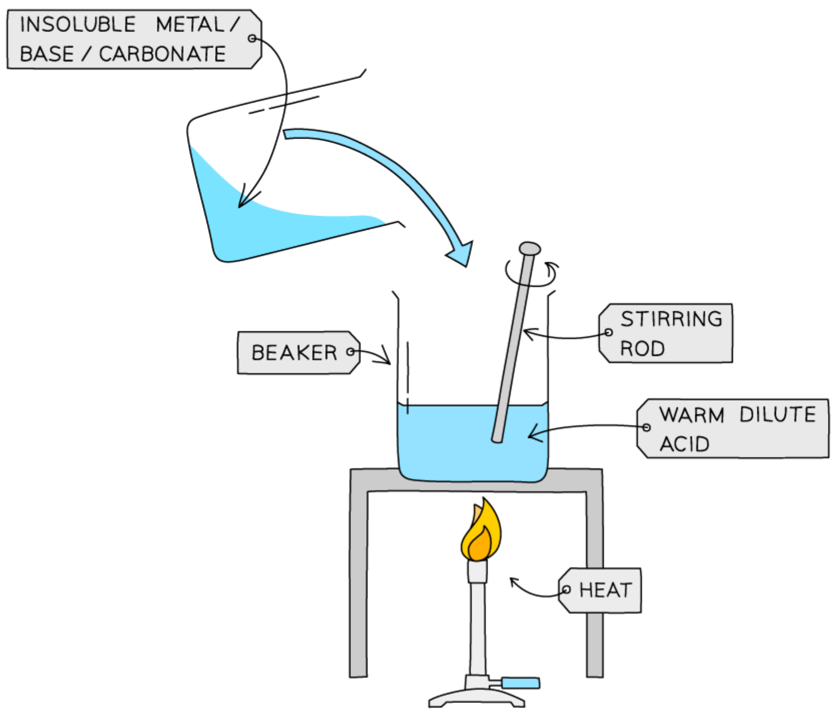
Preparing soluble salts

Method A: adding acid to a solid metal, insoluble base or insoluble carbonate

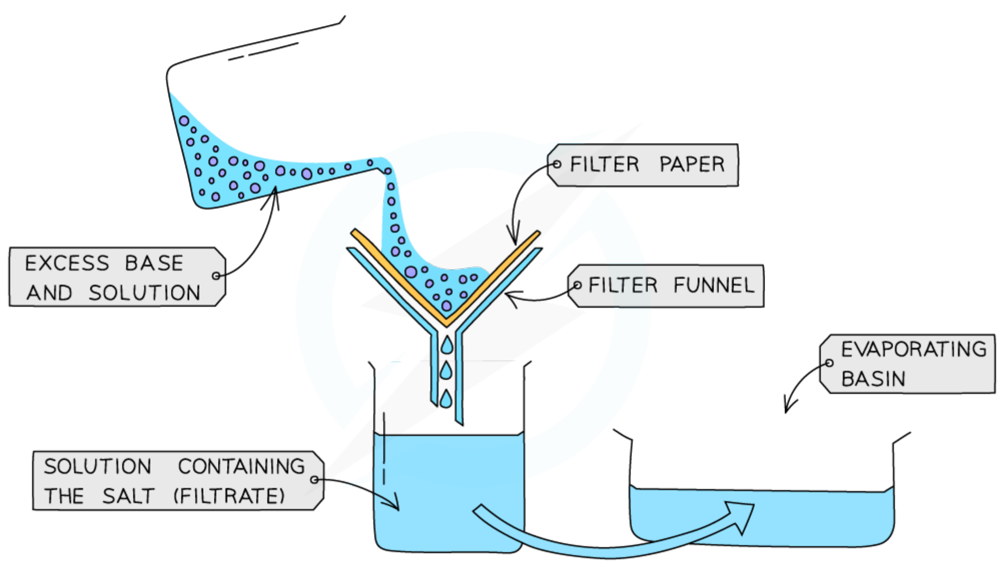
© 2015-2021 Save My Exams, Ltd. · Revision Notes, Topic Questions, Past Papers

YOUR NOTES
↓

1 HEAT ACID UNTIL WARM, THEN ADD METAL / BASE / CARBONATE, STIRRING CONSTANTLY UNTIL IT STOPS DISAPPEARING



2 FILTER MIXTURE TO REMOVE EXCESS BASE, TRANSFER SOLUTION TO EVAPORATING BASIN.



© 2015-2021 Save My Exams, Ltd. · Revision Notes, Topic Questions, Past Papers

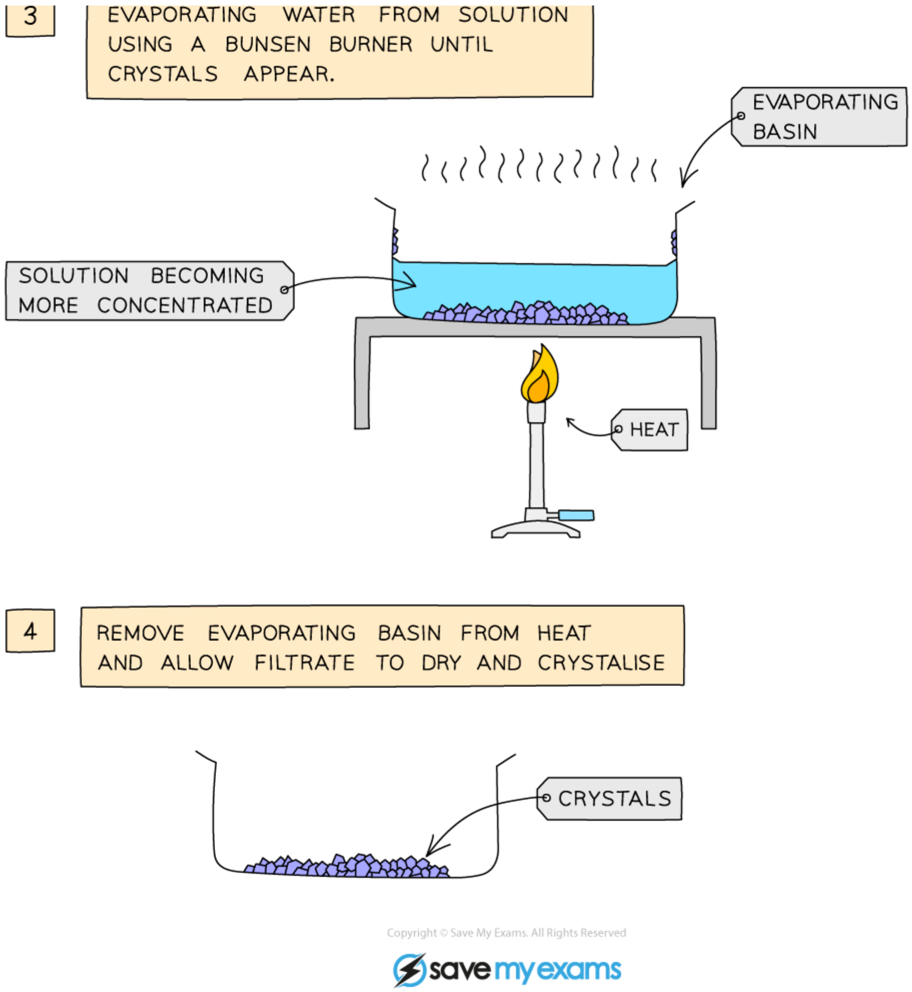


Diagram showing the preparation of soluble salts

Method:

- Add dilute acid into a beaker and heat using a bunsen burner flame
- Add the insoluble metal, base or carbonate, a little at a time, to the warm dilute acid and stir until the base is in excess (i.e. until the base stops disappearing and a suspension of the base forms in the acid)
- Filter the mixture into an evaporating basin to remove the excess base
- Heat the solution to evaporate water and to make the solution saturated. Check the solution is saturated by dipping a cold, glass rod into the solution and seeing if crystals form on the end
- Leave the filtrate in a warm place to dry and crystallize
- Decant excess solution and allow crystals to dry or blot to dry with filter paper

Example: preparation of pure, hydrated copper(II) sulfate crystals using method A

Acid = dilute sulfuric acid

© 2015-2021 Save My Exams, Ltd. · Revision Notes, Topic Questions, Past Papers

Insoluble base = copper(II) oxide

YOUR NOTES
⬇

Method:

- Add dilute sulfuric acid into a beaker and heat using a bunsen burner flame
- Add copper(II) oxide (insoluble base), a little at a time to the warm dilute sulfuric acid and stir until the copper (II) oxide is in excess (stops disappearing)
- Filter the mixture into an evaporating basin to remove the excess copper(II) oxide
- Leave the filtrate in a warm place to dry and crystallize
- Decant excess solution
- Blot crystals dry with filter paper

Equation of reaction:

copper(II) oxide + sulfuric acid → copper(II) sulphate + water

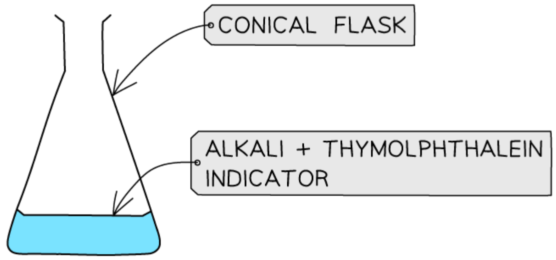
$$CuO\ (s) + H_2SO_4\ (aq) \rightarrow CuSO_4\ (aq) + H_2O\ (l)$$

Method B: reacting a dilute acid and alkali (soluble base)

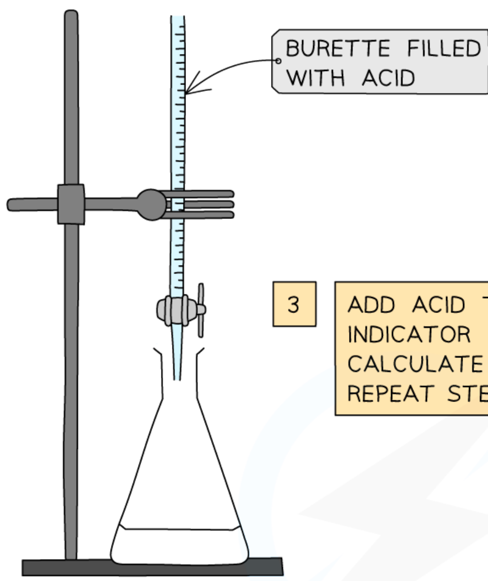
© 2015-2021 Save My Exams, Ltd. · Revision Notes, Topic Questions, Past Papers

YOUR NOTES
⬇

1 ADD ALKALI + INDICATOR TO CONICAL FLASK USING A PIPETTE

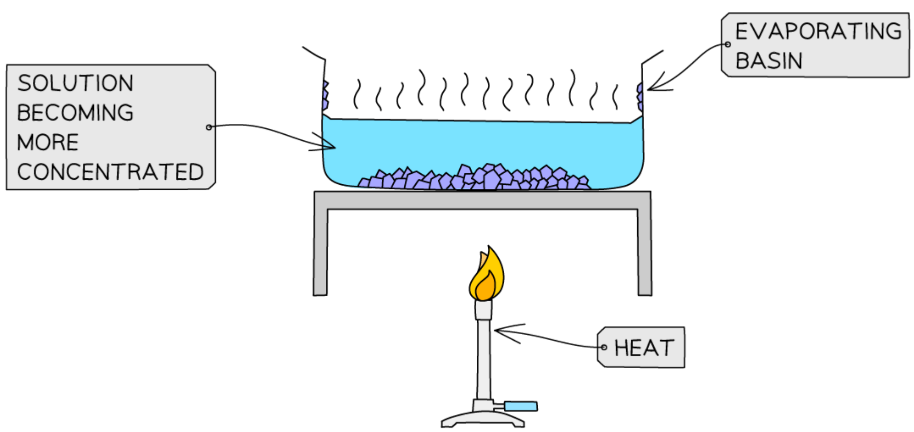


2 ADD ACID TO BURETTE, NOTING THE STARTING VOLUME



3 ADD ACID TO ALKALI SLOWLY UNTIL INDICATOR CHANGES COLOUR. CALCULATE VOLUME OF ACID ADDED. REPEAT STEPS 1–3 WITHOUT INDICATOR

4 TRANSFER SOLUTION TO AN EVAPORATING BASIN, HEAT TO PARTIALLY EVAPORATE WATER



5 REMOVE EVAPORATING BASIN FROM HEAT AND ALLOW FILTRATE TO DRY AND CRYSTALISE

© 2015-2021 Save My Exams, Ltd. · Revision Notes, Topic Questions, Past Papers

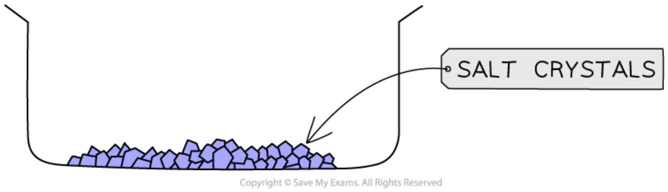


Diagram showing the apparatus needed to prepare a salt by titration

Method:

- Use a pipette to measure the alkali into a conical flask and add a few drops of indicator (thymolphthalein 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 the appropriate colour
- Note and record the final volume of acid in the 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 the resulting solution in an evaporating basin to partially evaporate, leaving a saturated solution (crystals just forming on the sides of the basin or on a glass rod dipped in and then removed)
- Leave to crystallise, decant excess solution and allow crystals to dry

YOUR NOTES
⬇

© 2015-2021 Save My Exams, Ltd. · Revision Notes, Topic Questions, Past Papers

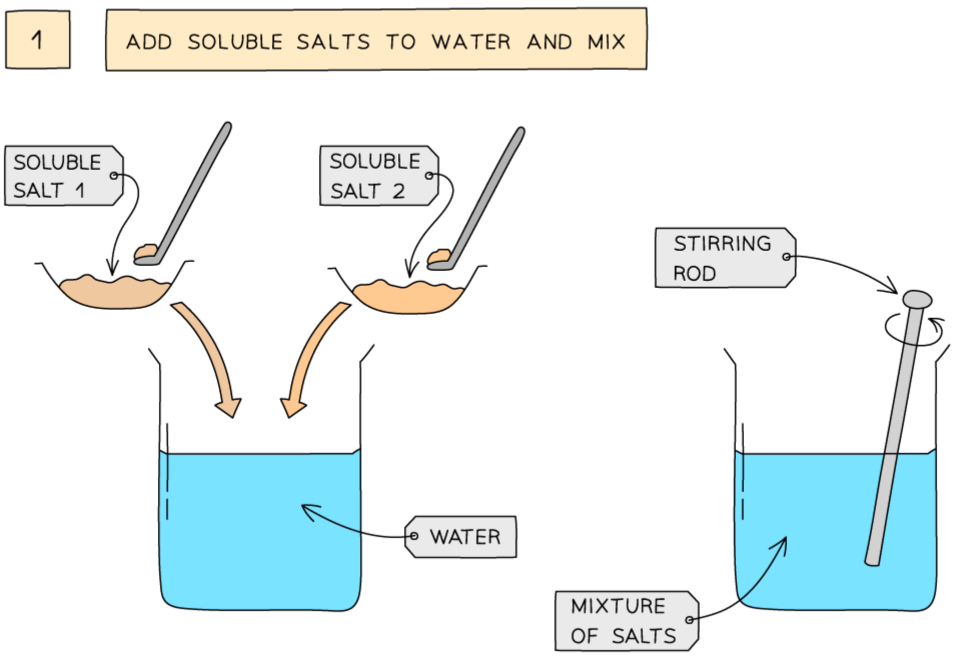
YOUR NOTES
↓

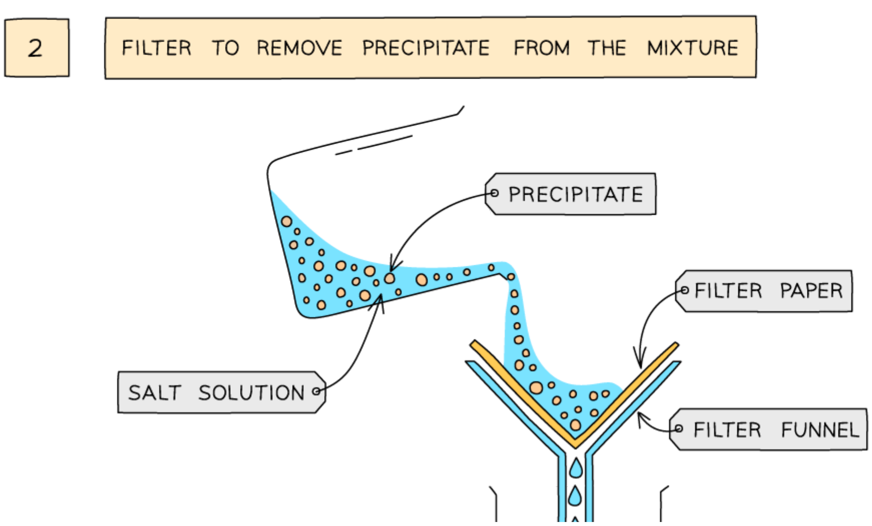
Preparing Insoluble Salts

EXTENDED

- Insoluble salts can be prepared using a precipitation reaction
- The solid salt obtained is the precipitate, thus in order to successfully use this method the solid salt being formed must be insoluble in water, and the reactants must be soluble

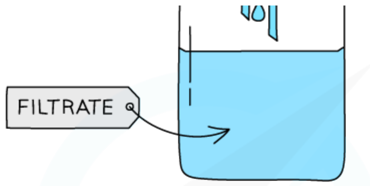
Using two soluble reactants



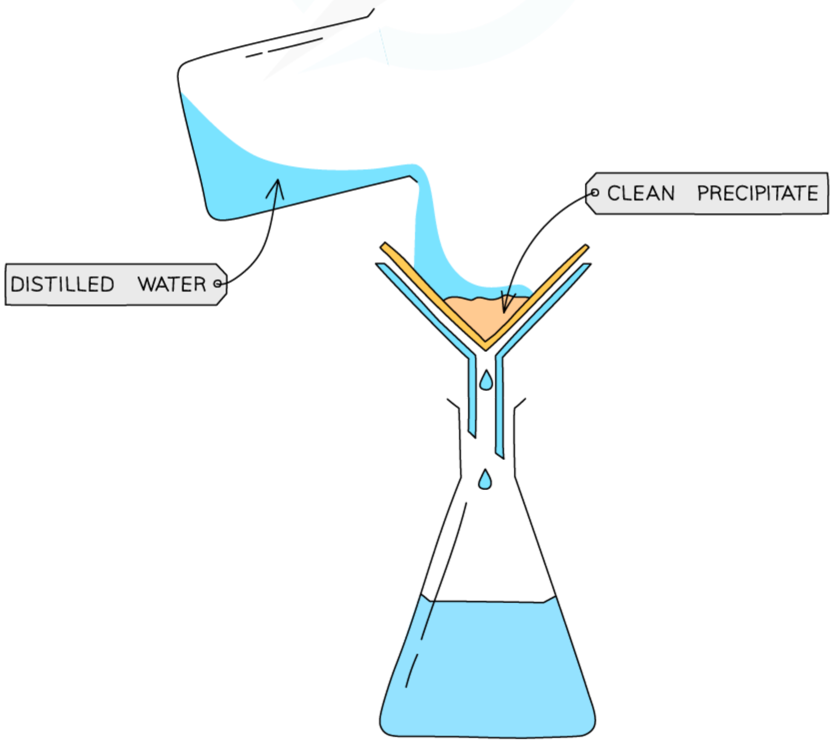


© 2015-2021 [Save My Exams, Ltd.](https://savemyexams.co.uk) · Revision Notes, Topic Questions, Past Papers

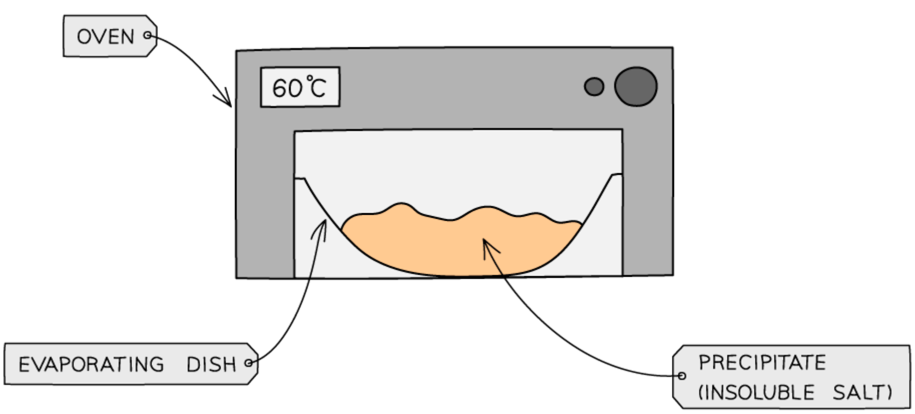
YOUR NOTES



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



4 DRY THE PRECIPITATE (INSOLUBLE SALT) IN AN OVEN



© 2015-2021 Save My Exams, Ltd. · Revision Notes, Topic Questions, Past Papers

Copyright © Save My Exams. All Rights Reserved

Diagram showing the filtration of the mixture to remove the precipitate

YOUR NOTES

Method:

- Dissolve soluble salts in water and mix together using a stirring rod in a beaker
- Filter to remove precipitate from mixture
- Wash filtrate with distilled water to remove traces of other solutions
- Leave in an oven to dry

Example: Preparation of pure, dry lead(II) sulfate crystals using a precipitation reaction

Soluble Salt 1 = lead(II) nitrate

Soluble Salt 2 = potassium sulfate

Method:

- Dissolve lead(II) nitrate and potassium sulfate in water and mix together using a stirring rod in a beaker
- Filter to remove precipitate from mixture
- Wash precipitate with distilled water to remove traces of potassium nitrate solution
- Leave in an oven to dry

Equation of reaction:

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

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

© 2015-2021 Save My Exams, Ltd. · Revision Notes, Topic Questions, Past Papers

7.2.3 Solubility Rules

YOUR NOTES
↓

Solubility Rules

- Salts are prepared by different methods, depending on whether the salt is soluble or insoluble so it is important to know the solubility of salts

Solubility of the common salts

Salts	Soluble	Insoluble
Sodium, potassium and ammonium	All	None
Nitrates	All	None
Chlorides	Most are soluble	Silver and lead(II)
Sulfates	Most are soluble	Barium, calcium and lead(II)
Carbonates	Carbonates of sodium, potassium and ammonium	Most are insoluble
Hydroxides	Hydroxides of sodium, potassium, ammonium and calcium (calcium hydroxide is partially soluble	Most are insoluble

Copyright © Save My Exams. All Rights Reserved

7.2.4 Hydrated & Anhydrous Salts

YOUR NOTES
↓

Hydrated & Anhydrous Salts

- When salts are being prepared, some water can be retained within the structure of the salt during the crystallisation process
- This affects the crystal's shape and colour
- Salts that contain water within their structure are called hydrated salts
- Anhydrous salts are those that contain no water in their structure
- A common example is copper(II) sulfate which crystallises forming the salt hydrated copper(II) sulfate, which is blue
- When it is heated, the water from its structure is removed, forming anhydrous copper(II) sulfate, which is white
- The hydrated salt has been **dehydrated** to form the anhydrous salt
- This reaction can be reversed by adding water to anhydrous copper(II) sulfate:

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

© 2015-2021 Save My Exams, Ltd. · Revision Notes, Topic Questions, Past Papers

Water of Crystallisation

YOUR NOTES

EXTENDED

- Water molecules included in the structure of some salts during the crystallisation process are known as **water of crystallisation**
- A compound that contains water of crystallisation is called a **hydrated compound**
- When writing the chemical formula of hydrated compounds, the water of crystallisation is separated from the main formula by a **dot**
  - E.g. hydrated copper(II) sulfate is $CuSO_4 \cdot 5H_2O$
  - Hydrated cobalt(II) chloride is $CoCl_2 \cdot 6H_2O$
- The formula shows the number of moles of water contained within one mole of the hydrated salt
  - E.g. hydrated copper(II) sulfate, $CuSO_4 \cdot 5H_2O$, contains 5 moles of water in 1 mole of hydrated salt
- A compound which doesn't contain water of crystallisation is called an **anhydrous compound**
  - E.g. anhydrous copper(II) sulfate is $CuSO_4$
  - Anhydrous cobalt(II) chloride is $CoCl_2$
- The conversion of anhydrous compounds to hydrated compounds is **reversible** by heating the hydrated salt:
  - Anhydrous to hydrated salt:
    - $\mathbf{CuSO_4 + 5H_2O \rightarrow CuSO_4 \cdot 5H_2O}$
  - Hydrated to anhydrous salt (by heating):
    - $\mathbf{CuSO_4 \cdot 5H_2O \rightarrow CuSO_4 + 5H_2O}$

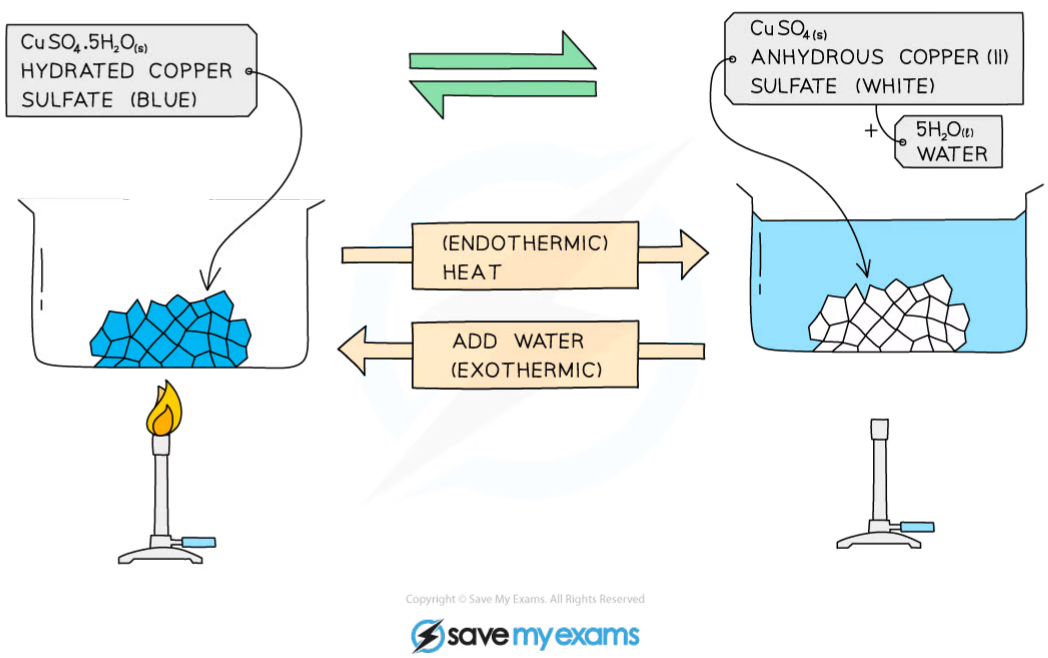


Diagram showing the dehydration of hydrated copper II) sulfate

© 2015-2021 Save My Exams, Ltd. · Revision Notes, Topic Questions, Past Papers

© 2015-2021 Save My Exams, Ltd. · Revision Notes, Topic Questions, Past Papers