AISS Chemistry Unit
Chapter 12: Salts

Name: ……..( ) Class: …….. Date: …….

At the end of this section, you should be able to:
- *define a salt*
- *state common reactions used to prepare a salt*
- *describe the general rules for the solubility of common salts in water*

12.1 What is a Salt?

Based on your knowledge on the chemical reactions in the chapter of acids and bases, state the possible reactions that would allow you to prepare a salt.

- Acid + metal → salt + hydrogen gas
  2HCl + Zn → $ZnCl_2$ + H_2

- Acid + base → salt + water
  H_2SO_4 + $2NH_3$ → $(NH_4)_2SO_4$

- Acid + metal carbonate → salt + carbon dioxide + water
  $2HNO_3$ + $CaCO_3$ → $Ca(NO_3)_2$ + CO_2 + H_2O

A **salt** is a compound formed when the **hydrogen ion in an acid is replaced by a metallic ion or an ammonium ion.**

Reactions in Which Salts Can Be Made

Possible reactants		**Salt formed**
Metal/ Carbonate (for the cation)	**Acid (for the anion)**	
zinc (Zn)	hydrochloric acid (HCl)	Zinc chloride ($ZnCl_2$)
copper(II) carbonate ($CuCO_3$)	Nitric acid (HNO_3)	copper(II) nitrate $Cu(NO_3)_2$

Magnesium (Mg) / Magnesium carbonate ($MgCO_3$)	sulfuric acid (H_2SO_4)	magnesium sulfate ($MgSO_4$)
aqueous ammonia NH_3(aq)	sulfuric acid (H_2SO_4)	Ammonium chloride NH_4Cl
potassium hydroxide (KOH)	Phosphoric acid (H_3PO_4)	potassium phosphate (K_3PO_4)

Checkpoint 1

State possible starting reactants for the following salts based on the above chemical acid/base reactions.

No	Reactants	Products (Salts)
1	Copper(II) oxide/hydroxide/carbonate + nitric acid	copper(II) nitrate
2	Sodium oxide/hydroxide/carbonate + hydrochloric acid	sodium chloride
3	Potassium oxide/hydroxide/carbonate + sulfuric acid	potassium sulfate
4	Zinc/zinc oxide/hydroxide/carbonate + hydrochloric acid	zinc chloride
5	magnesium/magnesium oxide/hydroxide/carbonate + nitric acid	magnesium nitrate

12.2 General rules for the solubility of common salts in water

Do ionic compounds dissolve in water?

Yes. *Water molecules are attracted to ions. This weakens the electrostatic forces between the ions. The ions are pulled from the lattice structure and the compound dissolves to form aqueous solution.*

However, solubility in water is dependent on temperature and nature of solute as well.

The table below summarises the general rules for solubility of the common salts in water at room temperature.

Soluble salts		Insoluble salts
All **s**odium, **p**otassium, **a**mmonium salts (**SPA**)		
All nitrates		
All **C**hlorides *(**C**lown **L**augh & **S**mile)*	except ⇒	**L**ead chloride, $PbCl_2$ **S**ilver chloride, $AgCl$
All **S**ulfates *(**S**uper **B**ig **C**hicken **L**eg)*	except ⇒	**B**arium sulfate, $BaSO_4$ **C**alcium sulfate, $CaSO_4$ **L**ead sulfate, $PbSO_4$
Sodium carbonate, Na_2CO_3 Potassium carbonate, K_2CO_3 Ammonium carbonate, $(NH_4)_2CO_3$	except ⇐	All carbonates
Sodium hydroxide, NaOH Potassium hydroxide, KOH Ammonium hydroxide, NH_4OH	except ⇐	All hydroxides Calcium hydroxide, $Ca(OH)_2$ is slightly soluble
Sodium oxide, Na_2O Potassium oxide, K_2O	except ⇐	All oxides Calcium oxide, CaO is slightly soluble

Not salts, these are bases.

The solubility of a salt must be determined before we can choose a suitable method for preparing the salt.

Checkpoint 2

State the solubility of the following salts/bases.

No	Salts	Solubility (Soluble/Insoluble)
1	copper(II) nitrate	Soluble
2	sodium chloride	Soluble
3	barium sulfate	Insoluble
4	zinc chloride	Soluble
5	magnesium hydroxide	Insoluble

12.3 Water of Crystallisation

- Salts crystals are often formed by crystallisation from aqueous solutions, and thus often have water molecules bonded to them. This water is known as water of crystallisation.
- Salts that contain water of crystallisation are known as hydrated salts.
- Heating a hydrated salt drives away the water, leaving the anhydrous salt.

$$\text{hydrated salt} \xrightarrow{\text{heat}} \text{anhydrous salt} + \text{water}$$

Examples of salts containing water of crystallisation

(1) **Copper(II) sulfate**

$CuSO_4.5H_2O(s)$	⟶	$CuSO_4(s)$	+	$5H_2O(g)$
hydrated copper(II) sulfate, also known as copper(II) sulfate pentahydrate		anhydrous copper(II) sulfate		water vapour
blue crystalline powder		white powder		water of crystallisation lost as water vapour

(2) **Cobalt(II) chloride**

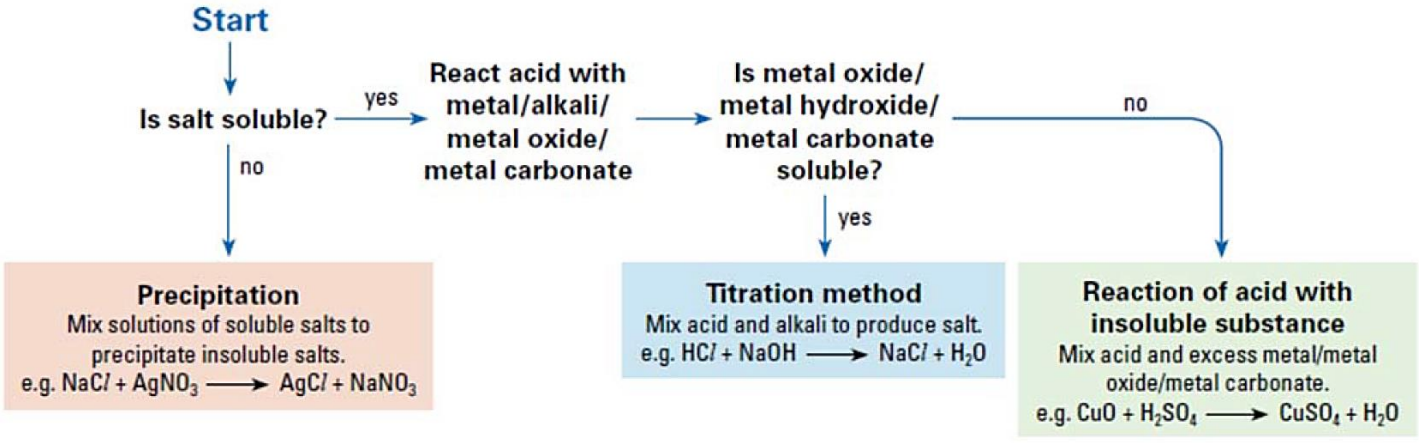
$CoCl_2.6H_2O(s)$	⟶	$CoCl_2(s)$	+	$6H_2O(g)$
hydrated cobalt(II) chloride, also known as cobalt(II) chloride hexahydrate		anhydrous cobalt(II) chloride		water vapour
pink		blue		water of crystallisation lost as water vapour

12.4 Preparation of Salts

Recall: Solubility table

The method of preparation depends on three factors:

(a) Solubility of the salt
(b) Reactivity of the metal
(c) Chemicals being reacted together to make the salt



Flowchart for salt preparation method

At the end of this section 2, you should be able to:
- *Correlate practical experiment to the chemical equation and state symbols*
- *Understand the rationale of the steps to prepare soluble and insoluble salts*
- *Suggest suitable separation techniques to obtain soluble and insoluble salts*
- *Suggest suitable reactants to obtain soluble and insoluble salt*

12.4.1 Preparation of Soluble Salts

- The following are different ways to prepare a **soluble** salt.

1. Reaction between an acid and excess of an insoluble solid:
a) Reacting an acid with an insoluble **carbonate**
b) Reacting an acid with an insoluble **base**
c) Reacting an acid with a **metal**
- Never use reactive metals, e.g. sodium, potassium, as the reaction is highly dangerous.
- Never use unreactive metals e.g. copper, gold, silver, platinum
- Only moderately reactive metals such as magnesium, zinc are suitable for this method.

2. Titration
a) Reacting an acid with an **alkali** (soluble base)
b) Reacting an acid with a **soluble carbonate** (SPA carbonate)

(1) Reacting an acid with an insoluble carbonate

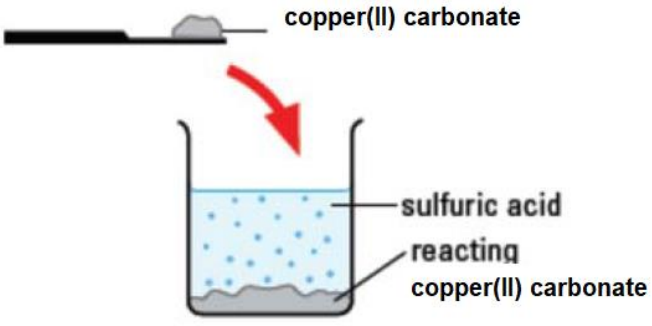
E.g. Preparation of pure and dry copper(II) sulfate, $CuSO_4$ crystals

Step 1 Identify solubility of salt	 Copper(II) sulfate, $CuSO_4$, is soluble in water (Not SPA)
Step 2 Identify the appropriate method	Thought process: 1. How do I decide the acid to be used? Acid to be used is decided by identity of the anion of the salt. Acid → sulfuric acid 2. Why can't a metal be used to prepare this salt? Copper is an unreactive metal, so copper will not react with acid to form salt. Other unreactive metals include, silver, gold and platinum

	3. Can an insoluble base be used to prepare this salt? If yes, give the possible insoluble bases. Copper(II) oxide Copper(II) hydroxide Copper(II) carbonate
	Write the balanced chemical equations, with state symbols for the possible reactants used with acid to prepare copper(II) sulfate. (1) $CuO(s) + H_2SO_4(aq) \rightarrow CuSO_4(aq) + H_2O(l)$ (2) $CuCO_3(s) + H_2SO_4(aq) \rightarrow CuSO_4(aq) + H_2O(l) + CO_2(g)$ (3) $Cu(OH)_2 + H_2SO_4(aq) \rightarrow CuSO_4(aq) + 2H_2O(l)$
Step 3 **Prepare the salt by following the steps**	

Teacher can choose to do demonstration or play video in class.
→copper(II) carbonate + sulfuric acid

Step 1:
Fill half a beaker with dilute sulfuric acid. With constant stirring, add **solid** copper(II) carbonate until **no more copper(II) carbonate can be dissolved**.



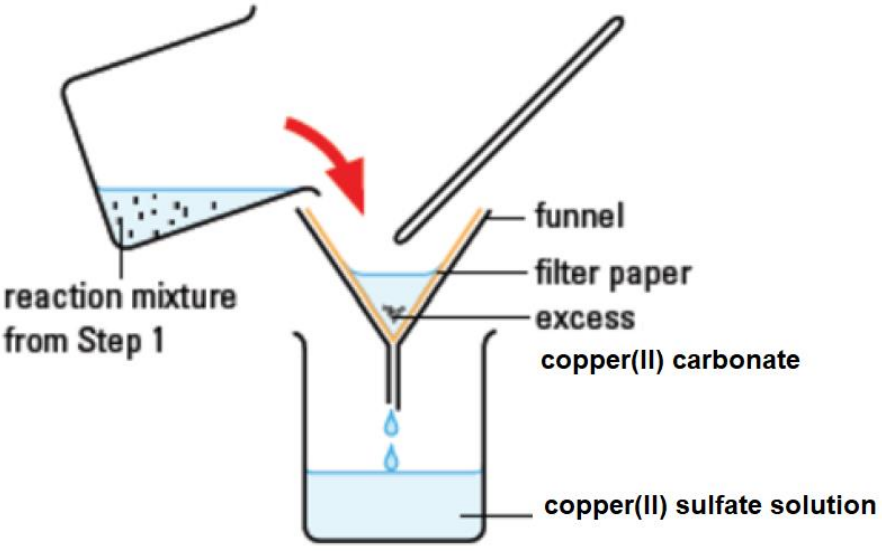
See	Think
Before addition of chemicals: Colour and state of (Appearance) • Sulfuric acid: colourless solution • Copper(II) carbonate: green solid	Copper is a transition metal, forming coloured compounds

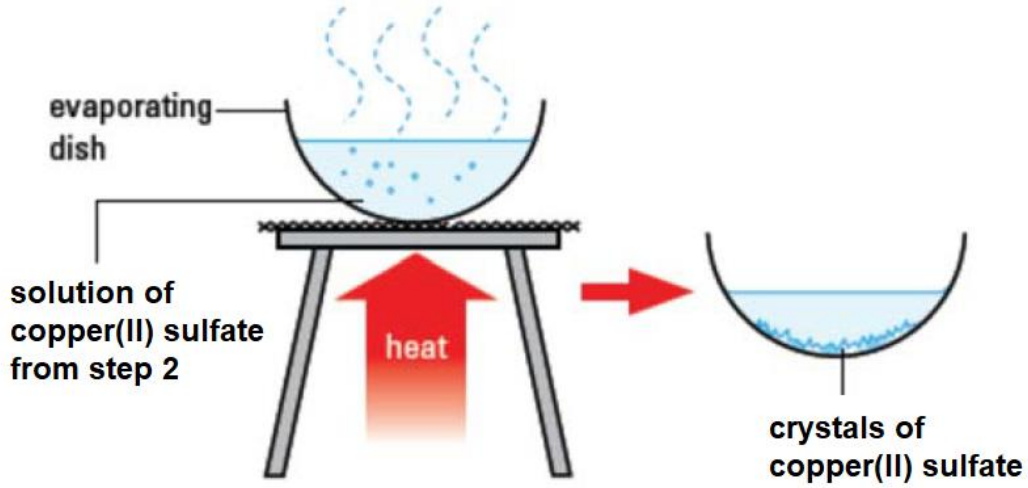
During the addition of chemicals, what do you observed?	
1. effervescence is observed	• Production of CO_2
2. Green solid dissolved as it is added	• All copper(II) carbonate reacts with the sulfuric acid to form copper(II) sulfate
3. Colourless solution turns blue	• Colour of copper(II) sulfate solution
4. Colour of blue litmus paper (turns red)	• Acid still in excess
Excess insoluble solid added 1. Effervescence	• Production of CO_2 as reaction is still going on
2. Green solid could no longer dissolve → implies excess insoluble solid	• Excess copper(II) carbonate is added. No more acid is present.
3. Colour of blue litmus paper	• Blue litmus remains blue. No acid present in the flask/beaker

Step 2:

Filter the mixture to remove **excess** unreacted copper(II) carbonate.

Residue is excess copper(II) carbonate and filtrate is copper(II) sulfate solution.





Step 3: **Heat** the filtrate until the solution is saturated.

Step 4: Allow the saturated solution to **cool** so that copper(II) sulfate salt can crystallise.

Step 5: **Filter** to collect the crystals. **Wash** the crystal with small volume of cold deionised water to remove impurities.

Step 6: **Dry** the crystals between a few **sheets of filter paper.**

[The above method can be used when acid reacts with insoluble base and metal]

Checkpoint 3

Suggest the chemicals used to prepare the following soluble salts.
Write the chemical equation with state symbols for each reaction.

1 Zinc chloride

Chemicals: Hydrochloric acid and Zinc/zinc oxide/hydroxide/carbonate

Chemical Equation: $2HCl(aq) + ZnCO_3(s) \rightarrow ZnCl_2(aq) + CO_2(g) + H_2O(l)$

2 Silver nitrate

Chemicals: ntiric acid and silver oxide/hydroxide/carbonate

Chemical Equation: $2HNO_3(aq) + Ag_2CO_3(s) \rightarrow 2AgNO_3(aq) + CO_2(g) + H_2O(l)$

3 Magnesium chloride

Chemicals: Hydrochloric acid and magnesium/magnesium oxide/hydroxide/carbonate

Chemical Equation: $2HCl(aq) + Mg(s) \rightarrow MgCl_2(aq) + H_2(g)$

- Very reactive metals such as (Li, Na, K, Ca) react violently with acids, so the reaction is very dangerous.
- NaO, NaOH, $NaCO_3$ are all soluble salts. Can you still use method 1?

(2) Reacting an acid with alkali via titration method

- We can prepare a soluble salt by reacting with an acid with an alkali. However, because both reactants are usually colourless, and the products are also colourless and soluble in water, we have no way of determining when all the acid has completely reached with all the alkali. Hence. We use titration.

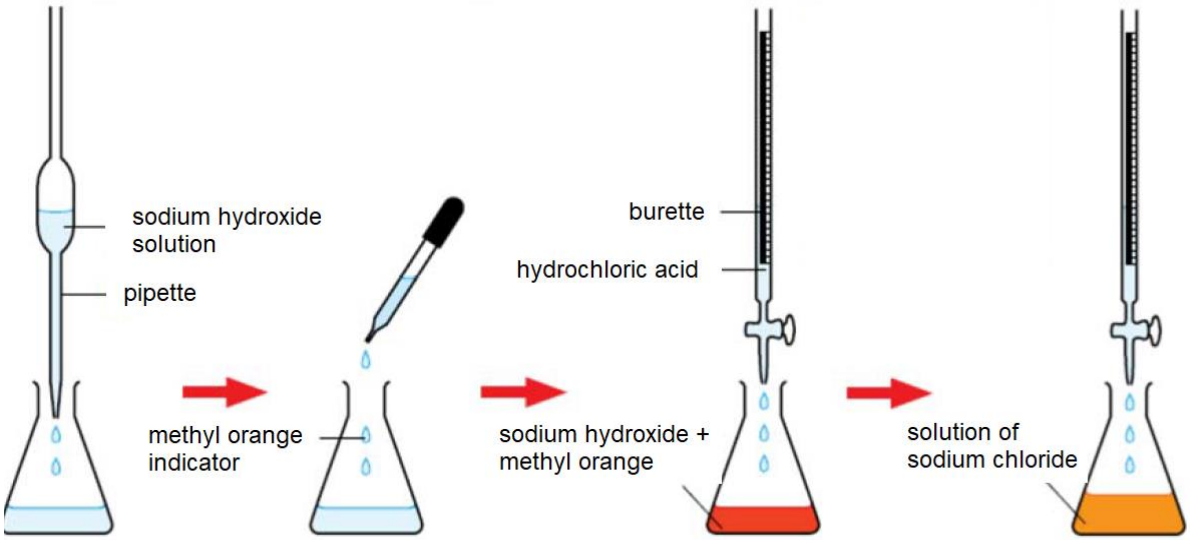
Preparation of pure and dry sodium chloride, a soluble salt.

Step 1 Identify solubility of salt	 sodium chloride, NaCl, is soluble in water
Step 2 Identify method of preparation	 NaCl contains the chloride anion → use hydrochloric acid What are possible reactants to react with hydrochloric acid? Sodium carbonate, sodium oxide, sodium hydroxide Why is titration the most appropriate method to prepare sodium sulfate? Sodium carbonate, sodium oxide, sodium hydroxide are soluble in water. There is no way of determining if excess solid is added or only salt remains at the end of reaction. Thus titration with an indicator will be able to determine when the reaction is complete.

Step 3

Summary of steps involved in preparation of <u>soluble salts</u>

Prepare the salt by following the steps.



Step 1: Fill up a burette with hydrochloric acid. Note the initial burette reading (V_1 cm^3).

Step 2: Pipette 25.0cm^3 of sodium hydroxide into a conical flask

Step 3: Add 2 drops of methyl orange to the conical flask with the sodium hydroxide solution. The solution will turn yellow.

Step 4: While swirling, add hydrochloric acid from the burette slowly until solution turns orange permanently. This is the end-point. Record the final burette reading (V_2 cm^3).

Step 5: Find the volume of hydrochloric acid required for complete neutralisation (V_2-V_1 cm^3)

Step 6: Pipette 25.0 cm^3 of sodium hydroxide into another conical flask. Add (V_2-V_1 cm^3) of hydrochloric acid from the burette. (Do not add indicator as it will make the salt impure)

Step 7: Heat the solution until it is saturated.

Step 8: Allow the saturated solution to cool so that sodium chloride salt can crystallise.

Step 9: Filter to collect the crystals. Wash the crystal with a little cold deionised water to remove impurities.

Step 10: Dry the crystals between a few sheets of filter paper.

Additional note:

Bromothymol blue (yellow(<6) --- green --- blue (>8)) ;

methyl orange (red(<3) --- orange --- yellow (>5))

Observation of the different colours of solution in a conical flask

See	Think
Before titration: Colour of solution in the conical flask with methyl orange indicator • yellow	• alkali has a pH above 7, methyl orange is yellow under basic conditions.
At endpoint: Colour of solution in the conical flask with methyl orange indicator • yellow to orange	• at the end point, all the alkali has been reacted by the amount of acid added. Only salt and water left in the flask.
Beyond endpoint: Colour of solution in the conical flask with methyl orange indicator • orange to red	• excess acid turning methyl orange to red.

Checkpoint 4

Suggest the chemicals used to prepare the following soluble salts.
Write the chemical equation with state symbols for each reaction.

1 Potassium nitrate

Method: Titration

Chemicals: Nitric acid and potassium hydroxide/carbonate

Chemical Equation: $KOH(aq) + HNO_3(aq) \rightarrow KNO_3(aq) + H_2O(l)$

2 Ammonium chloride

Method: Titration

Chemicals: Hydrochloric acid and aqueous ammonia

Chemical Equation: $HCl(aq) + NH_3(aq) \rightarrow NH_4Cl(aq)$

3 Iron(II) sulfate

Method: Reaction between acid and excess insoluble solid

Chemicals: sulfuric acid and iron hydroxide/carbonate/iron

Chemical Equation: $Fe(aq) + H_2SO_4(aq) \rightarrow FeSO_4(aq) + H_2(g)$

4 Barium chloride

Method: Reaction between acid and excess insoluble solid

Chemicals: Hydrochloric acid and barium/ barium oxide/ barium carbonate

Chemical Equation: $2HCl(aq) + BaCO_3(s) \rightarrow BaCl_2(aq) + H_2O(l) + CO_2(g)$

5 Sodium sulfate

Method: Titration

Chemicals: Sulfuric acid and sodium hydroxide/carbonate

Chemical Equation: $2NaOH(aq) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$

12.4.2 Preparation of Insoluble Salt

- Insoluble salts can be prepared by **precipitation**
- **Two aqueous solutions (soluble) salts** are used to produce insoluble salt, which can be obtained by **simple filtration method**.

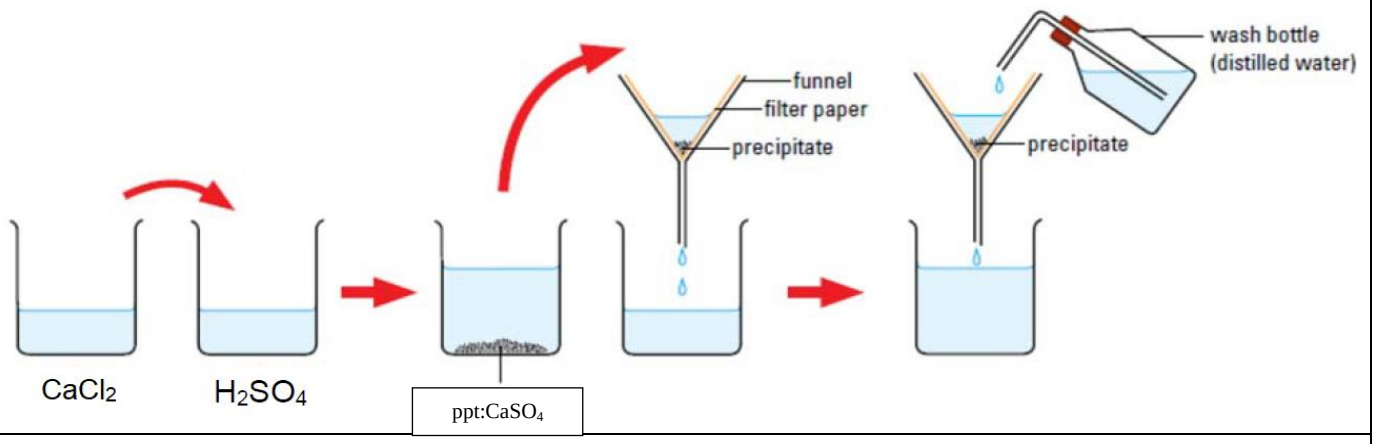
Teacher demo
Barium sulfate using barium nitrate and sulfuric acid

Preparation of insoluble salt, calcium sulfate.

Step 1 Identify solubility of salt	 calcium sulfate, $CaSO_4$, is insoluble in water
Step 2 Choose appropriate method	 Precipitation
Step 3 Identify the starting reactants Recall : Both reactants must be solutions. Either use (1) **Acid/alkali and soluble salt** or **(2) Both soluble salts** Q: instead of sulfuric acid, can sodium sulfate be used instead? A: Yes. For precipitation reaction, the starting reagents must contain the ion and be soluble.	Thought process: To ensure that **the product** obtained has o**nly one insoluble salt**, calcium sulfate, $CaSO_4$, the **other product must be soluble in water**. X (aq) + Y (aq) → $CaSO_4$ (s) + Z (aq) solution solution insoluble solution (1) To obtain calcium sulfate, $CaSO_4$, ➔ Use sulfuric acid, H_2SO_4 The other reactant can be soluble salt that contains calcium ion. Eg. $CaCl_2$, $Ca(NO_3)_2$, Hint: Nitrate salts is usually used as they are soluble H_2SO_4 (aq) + $CaCl_2$ (aq) → $CaSO_4$(s) + HCl(aq) (2) Both reactants are soluble salts. Any soluble salts with sulfate ions and calcium ions can be used. Na_2SO_4 (aq) + $CaCl_2$ (aq) → $CaSO_4$ (s) + NaCl(aq)

Step 4

Sets to prepare calcium sulfate, $CaSO_4$



(1) Add calcium chloride into 100 cm^3 of dilute sulfuric acid. Calcium sulfate ppt is formed.
(2) The resulting mixture is filtered to obtain the precipitate in the filter paper.
(3) The precipitate is washed with a small volume of cold deionised water to remove any impurities present.
(4) Dry the precipitate between pieces of filter paper.

Checkpoint 5

Suggest the two reactants that can be used to produce the following salts and write the balanced equations with state symbols.

No	Salts	Possible reactants
1	silver chloride	Silver nitrate + sodium chloride $AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$
2	lead(II) sulfate	Lead(II) nitrate + sodium sulfate $Pb(NO_3)_2(aq) + Na_2SO_4(aq) \rightarrow PbSO_4(s) + 2NaNO_3(aq)$
3	calcium carbonate	Calcium nitrate + potassium carbonate $Ca(NO_3)_2(aq) + K_2CO_3(aq) \rightarrow CaCO_3(s) + 2KNO_3(aq)$
4	barium sulfate	Barium nitrate + sulfuric acid $Ba(NO_3)_2(aq) + H_2SO_4(aq) \rightarrow BaSO_4(s) + 2HNO_3(aq)$
5	lead(II) chloride	Lead(II) nitrate + sodium chloride $Pb(NO_3)_2(aq) + 2NaCl(aq) \rightarrow PbCl_2(s) + 2NaCl(aq)$

Tip: keep to nitrate and SPA salt/ acid for precipitation reactions.

You may use the examples in checkpoint to teach ionic equation writing.

Example 1

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

$Ag^+(aq) + \cancel{NO_3^-(aq)} + \cancel{Na^+(aq)} + Cl^-(aq) \rightarrow AgCl(s) + \cancel{Na^+(aq)} + \cancel{NO_3^-(aq)}$

Final ionic equation: $Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$

Teacher note: Understanding the rationale for the method of salt preparation is useful for practical examination.

Questions

1. In the preparation of insoluble salt, two aqueous solutions (acid/alkali or soluble salts) are used to react with each other. Why must two solutions be used as the starting reactants?

   - The salt produced is insoluble in water.
   - Using two solutions as the starting reactants allow the resulting insoluble salt to be separated from the soluble reactants and by-product by filtration.
   - The insoluble salt wont be able to dissolves to give the ions for precipitation to occur.

2. Why can't an insoluble salt be used as one of the reactants?

   The insoluble salt cannot dissolves in water to produce the necessary ions in water for precipitation to take place.

3. Why is the precipitate rinsed with cold distilled water instead of warm distilled water?

   - In general, precipitate formed is rinsed with cold distilled water instead of warm distilled water to prevent any possible dissolving of salt.
   - Note: for insoluble salt in precipitation, the dissolving of salt by rinsing is minimal due to its insoluble nature.

4. Why can metal/metal carbonate/metal oxide react with acid to form the insoluble salt?

   E.g. Preparation of calcium sulfate
   Why cant we use Calcium with sulfuric acid to produce calcium sulfate?
   The amt of calcium sulfate formed will be very little. This is because an insoluble layer of calcium sulfate will coat the calcium metal, preventing any further reaction.

Checkpoint 6

1. Which salt is best prepared by precipitation?

   A sodium sulfate
   B ammonium carbonate
   C zinc carbonate
   D potassium chloride

   C

2. Lead(II) chloride is insoluble in water. It can be prepared by reacting dilute hydrochloric acid with aqueous lead(II) nitrate.
   Which step is involved in the preparation of lead(II) chloride?

   A Titrate dilute hydrochloric acid against aqueous lead(II) nitrate.
   B Evaporate the filtrate until crystals of lead(II) chloride form.
   C Collect the precipitate of lead(II) chloride by filtration
   D Add dilute hydrochloric acid until no more gas is evolved.

   C

3. Which of the following methods produces the highest yield of calcium sulfate from calcium carbonate?

   A Addition of sulfuric acid
   B Addition of aqueous sodium sulfate
   C Addition of hydrochloric acid, followed by sulfuric acid
   D Addition of hydrochloric acid, followed by barium sulfate

   C

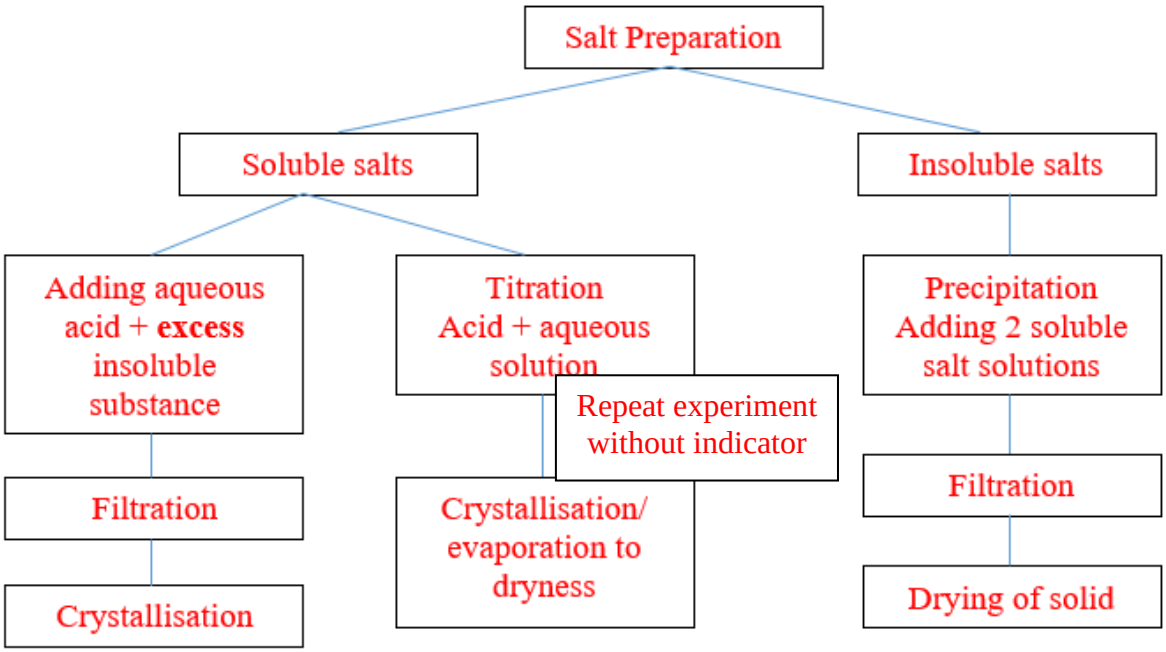
Recap:

Names of the three methods of salt preparation

Type of salt	**Examples of salts**	**Name of method**
Insoluble salts	$BaSO_4$. AgCl	Precipitation method
Soluble non-SPA salts	$MgCl_2$, $Zn(NO_3)_2$	Reaction between acid and insoluble solid (base/carbonate/metal)
Soluble SPA salts	NaCl, KNO_3	Titration method

Let us consolidate!

Create a flowchart to summarise what you have learnt about salt preparation. Add in details to make it comprehensive.



*when unsure, do crystallisation so as to get the hydrated salts.

Practice questions

Level 1

1. Which salt is best prepared by titration?

   A Barium sulfate
   B Calcium carbonate
   C Lead (II) nitrate
   D Potassium chloride

D
Titration is used to prepare salts of sodium, potassium and ammonium. This is due to the soluble nature of their carbonates, metal oxides and metal hydroxides.

2. Name the salt formed, if any, in each of the following reactions.
   State whether the salt formed is soluble in water.

Reactants	Name of salt formed	Solubility of salt formed (soluble / insoluble in water)
(a) zinc and nitric acid	Zinc nitrate	Soluble
(b) magnesium oxide and hydrochloric acid	Magnesium chloride	Soluble
(c) lead(II) nitrate and sodium sulfate	Lead(II) sulfate Sodium nitrate	Insoluble Soluble
(d) sodium hydroxide and sulfuric acid	Sodium sulfate	Soluble

Level 2

3. Name the chemicals and method you would use to prepare each of the salts listed below. Give the balanced chemical equations with state symbols.

   (a) Lead(II) carbonate

   Method: Precipitation

   Chemicals: Lead(II) nitrate and sodium/potassium carbonate

   Chemical equation: $Pb(NO_3)_2(aq) + Na_2CO_3(aq) \rightarrow 2NaNO_3(aq) + PbCO_3(s)$

   (b) Copper(II) nitrate

   Method: Reaction between acid and excess insoluble solid

   Chemicals: Copper(II) oxide/carbonate and nitric acid

   Chemical equation: $CuO(s) + 2HNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + H_2O(l)$

(c) Lithium chloride

Method: Titration

Chemicals: Hydrochloric acid and lithium hydroxide/carbonate

Chemical equation: $LiOH(aq) + HCl(aq) \rightarrow LiCl(aq) + H_2O(l)$

Level 3

4. Calcium carbonate is insoluble in water.
Calcium sulfate and calcium hydroxide are both sparingly soluble in water.

Three students were asked to suggest a method for making calcium sulfate from calcium carbonate:

Method 1: Add dilute nitric acid to calcium carbonate, then add dilute sulfuric acid to the calcium nitrate solution formed.

Method 2: Add dilute sulfuric acid to calcium carbonate.

Method 3: Heat calcium carbonate strongly to give calcium oxide, then react the calcium oxide with dilute sulfuric acid.

(a) Which method is the most suitable for preparing calcium sulfate?
Write the equation(s) for any reaction(s) involved. Include state symbols.

……..

(b) Explain why the other two methods are less suitable.

……..

……..

……..

Answer: 1

Method 1. The reactions are:

$CaCO_3 (s) + 2 HNO_3(aq) \rightarrow Ca(NO_3)_2 (aq) + H_2O (l) + CO_2 (g)$

$Ca(NO_3)_2 (aq) + H_2SO_4 (aq) \rightarrow CaSO_4 (s) + 2HNO_3 (aq)$

In methods 2 and 3, dilute sulfuric acid and calcium carbonate or calcium oxide react to from an insoluble layer of calcium sulfate, which prevents further reaction.

In method 3, calcium carbonate has to be heated to a very high temperature before it decomposes. This makes the process difficult and time consuming.

Extension by teacher

Another student suggested the following.

Method 4: Heat calcium carbonate strongly to give calcium oxide. React the calcium oxide with water to form calcium hydroxide solution and then add dilute sulfuric acid.

Given that calcium sulfate formed is insoluble, suggest a reason why this method may not give a high yield of calcium sulfate.

Method 4 may not give a high yield of calcium sulfate.

Calcium oxide is sparingly soluble. Thus there will only little amt of aqueous calcium hydroxide to react with sulfuric acid to form calcium sulfate.

5. (a) Name the chemicals and outline the key steps in preparing a pure sample of magnesium sulfate crystals.

Thought process:
What kind of salt am I preparing? → $MgSO_4$; it is a soluble salt
What method can I use? → either method 1 or 2
Possible ans:
Since the salt contains SO_4^{2-} ion → I must use sulfuric acid.
Reactants can be: sulfuric acid + Mg / Sulfuric acid + MgO / sulfuric acid + $Mg(OH)_2$ / sulfuric acid + $MgCO_3$
Mg is a metal and MgO, $Mg(OH)_2$ / $MgCO_3$ are all insoluble salts. So the reaction is between an aq acid and a solid → method 1.

Method: Reaction between acid and excess insoluble solid

Starting reagents: Sulfuric acid and magnesium/magnesium oxide/magnesium carbonate

Chemical equation: $H_2SO_4(aq) + Mg(s) \rightarrow MgSO_4(aq) + H_2(g)$

Procedure:

Step1: Fill half a beaker with dilute sulfuric acid. With constant stirring, add **solid** magnesium carbonate until **no more magnesium carbonate can be dissolved**.

Step2: **Filter** to remove **excess** unreacted magnesium carbonate. Residue is excess magnesium carbonate and filtrate is Magneisum sulfate solution.

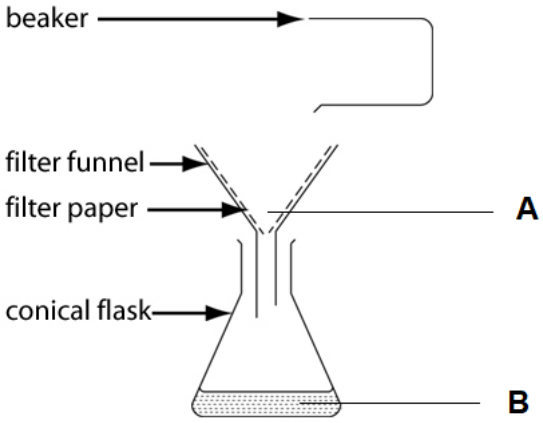
Step3: **Heat** the filtrate until it is **saturated**.

Step4: Allow the saturated solution to **cool** so that magneisum sulfate salt can crystallise.

Step5: **Filter** to collect the crystals. **Wash** the crystal with a **little cold** water to remove impurities.

Step6: **Dry** the crystals between a few sheets of filter paper. (unless qn ask for dry salt)

(b) With reference to the experimental procedure written above, the apparatus shown below can be used to filter the mixture. Identify residue **A** and solution **B**.



A : Magnesium/magnesium oxide/carbonate (Excess)

B: Magnesium sulfate solution

(c) Suggest, with a reason, why the crystals should be dried with filter paper and not by heating.

Magnesium sulfate crystals should not be dried by heating as it may lose its water of crystallisation and give anhydrous salts instead.

$MgSO_4 \cdot 7H_2O$.

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