Cambridge IGCSE · Chemistry 0620 · Notes

Preparation of salts

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Read first: **Acids, bases and oxides**, **Formulae and equations**, **Relative masses and the mole**

What you need to know

By the end of this topic you should be able to:

1. **Describe how to make a soluble salt**, then separate it and purify it, by reacting a dilute acid with:
   - (a) an **alkali, by titration**;
   - (b) an **excess of a metal**;
   - (c) an **excess of an insoluble base** (a metal oxide or hydroxide);
   - (d) an **excess of an insoluble carbonate**.
2. **Use the solubility rules**:
   - sodium, potassium and ammonium salts are all soluble;
   - nitrates are all soluble;
   - chlorides are soluble, except lead(II) chloride and silver chloride;
   - sulfates are soluble, except barium, calcium and lead(II) sulfate;
   - carbonates are insoluble, except sodium, potassium and ammonium carbonate;
   - hydroxides are insoluble, except sodium, potassium and ammonium hydroxide, and calcium hydroxide, which is partly soluble.
3. **Define a hydrated substance** as one that is chemically combined with water, and an **anhydrous substance** as one that contains no water.
4. **Extended only:** describe how to make an **insoluble salt by precipitation**.
5. **Extended only:** define **water of crystallisation** as the water molecules present in hydrated crystals, and know the examples $\mathrm{CuSO_4 \cdot 5H_2O}$ and $\mathrm{CoCl_2 \cdot 6H_2O}$.

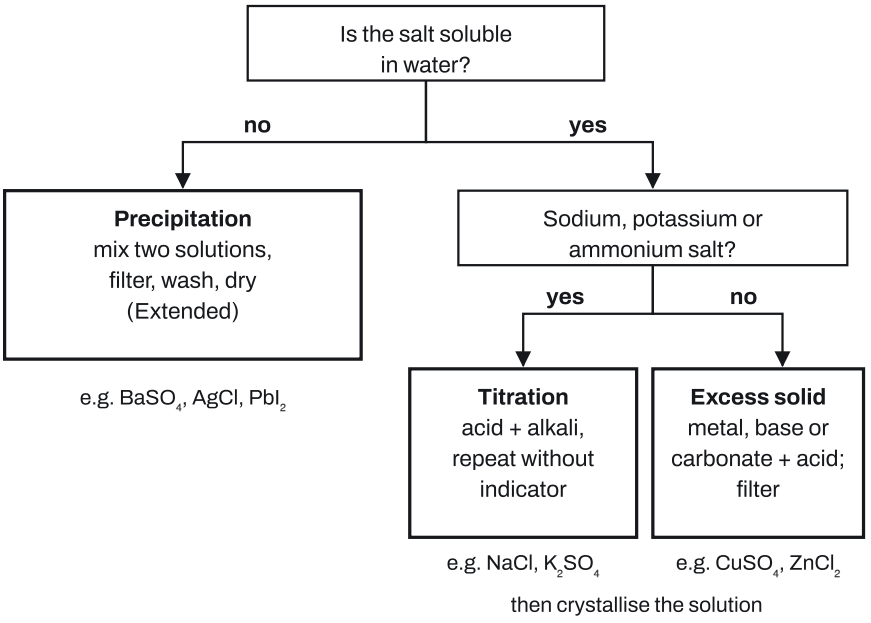
Understand it

What a salt is, and the first question to ask

A **salt** is the compound you get when the hydrogen of an acid is replaced by a metal (or by ammonium). The metal part comes from the other reactant; the negative ion comes from the acid: hydrochloric acid gives **chlorides**, sulfuric acid **sulfates** and nitric acid **nitrates**. The reactions themselves are the ones from Acids, bases and oxides:

- acid + metal → salt + hydrogen, for example $\mathrm{Zn + H_2SO_4 \rightarrow ZnSO_4 + H_2}$;
- acid + base → salt + water, for example $\mathrm{CuO + H_2SO_4 \rightarrow CuSO_4 + H_2O}$;
- acid + carbonate → salt + water + carbon dioxide, for example $\mathrm{ZnCO_3 + 2HCl \rightarrow ZnCl_2 + H_2O + CO_2}$.

To make a **pure, dry sample** of a salt you must also get rid of everything else in the flask. How you do that depends on one thing: **does the salt dissolve in water?** A soluble salt ends up dissolved, so you must crystallise it out of the solution. An insoluble salt appears as a solid, so you filter it off. The flowchart shows how to choose.



The solubility rules

You need to know which salts dissolve. Learn the rules as "everything in the first two rows dissolves; watch the exceptions in the rest".

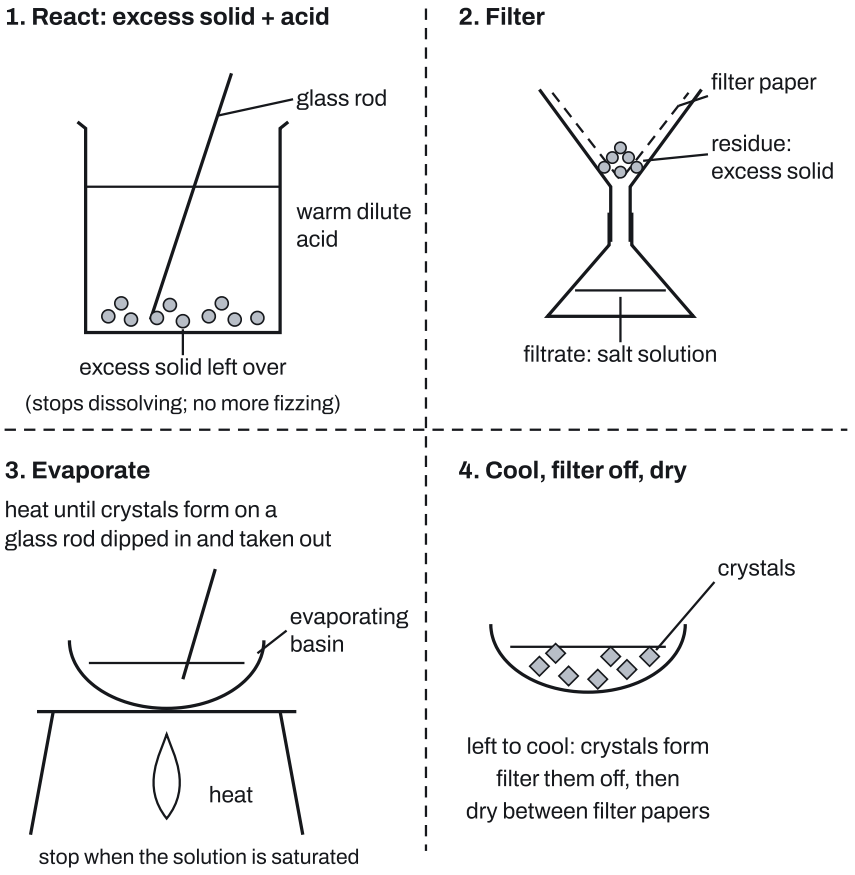
Salts of …	Soluble?
sodium, potassium, ammonium	**all soluble**
nitrates	**all soluble**
chlorides	soluble, **except** lead(II) chloride and silver chloride
sulfates	soluble, **except** barium, calcium and lead(II) sulfate
carbonates	**insoluble**, except sodium, potassium and ammonium carbonate
hydroxides	**insoluble**, except sodium, potassium and ammonium hydroxide; calcium hydroxide is **partly** soluble

The first two rows win over the others: ammonium carbonate and sodium sulfate are soluble; lead(II) nitrate and silver nitrate are soluble. This is why lead(II) nitrate, silver nitrate and barium nitrate (or barium chloride) are the usual sources of lead, silver and barium ions in solution.

Soluble salts: acid + an excess of an insoluble solid

This is the method for most soluble salts, such as copper(II) sulfate, zinc chloride or magnesium nitrate. The solid can be a **metal** (only a moderately reactive one: magnesium, zinc or iron), an **insoluble base** (a metal oxide or hydroxide) or an **insoluble carbonate**.

1. **React.** Warm the dilute acid (so the reaction is faster) and add the solid a little at a time, stirring, until **some solid is left over** and does not disappear. That leftover solid shows the solid is **in excess**. With a metal or a carbonate you will also see that the **fizzing stops**.
2. **Filter** off the excess solid. The solid left on the filter paper is the **residue**; the salt solution that passes through is the **filtrate**.
3. **Evaporate** the filtrate in an evaporating basin, heating gently until it is **saturated** (the **point of crystallisation**). To test, dip a glass rod in and take it out: if crystals form on the rod as it cools, stop heating.
4. **Crystallise.** Leave the solution to **cool**. Crystals form.
5. **Filter off the crystals** (or pick them out), and **dry them** by pressing them between filter papers, or leave them in a warm place.



Why use an excess of the solid? So that **all the acid is used up**. The acid is the **limiting reactant**. Any acid left over would stay in the filtrate and end up in the crystals, but excess solid is easy to remove because it doesn't dissolve: you just filter it off.

Why not boil it dry? Heating to dryness leaves any soluble impurities in the solid, can break the salt down, and drives off the **water of crystallisation**, so you get a powder instead of hydrated crystals.

Which metals don't work? Copper and silver are below hydrogen in the reactivity series and **don't react** with dilute acids, so copper(II) chloride cannot be made from copper + hydrochloric acid; use copper(II) oxide, hydroxide or carbonate instead. Sodium, potassium and calcium react dangerously fast, and they also react with the water in the acid, so an "excess" of them would leave the hydroxide in the solution.

Saturated solutions and why crystals form on cooling

A **saturated solution** is one that contains the **maximum concentration of solute dissolved in the solvent, at a stated temperature**. How much salt a solution can hold (its **solubility**) usually **increases as the temperature increases**. So a solution that is saturated while hot can't hold all of its salt once it cools down, and the extra salt comes out as crystals. Crystals form because **solubility decreases as the temperature decreases**, not because more water evaporates. That is also why cooling the solution further (in cold water or ice) gives more crystals.

Sodium, potassium and ammonium salts: titration

Sodium, potassium and ammonium bases and carbonates are **soluble**, so you can't add "an excess" and filter it off: the excess would stay dissolved and contaminate the salt. Instead you add **exactly** the right amount of alkali to the acid, found by **titration**.

1. Put a measured volume of the **alkali** (for example $25.0\ \mathrm{cm}^3$ of sodium hydroxide, from a pipette) into a conical flask and add a few drops of **indicator** (methyl orange or thymolphthalein).
2. Add the **acid from a burette**, swirling, until the indicator just changes colour: the **end-point**. Methyl orange goes from **yellow** (alkali) **to orange** at the end-point (red if you overshoot); thymolphthalein goes from **blue to colourless**. Note the volume of acid used.
3. **Repeat with the same volumes but without the indicator**, so the salt solution is not contaminated by it.
4. **Crystallise** as before: heat to the point of crystallisation, cool, filter off the crystals and dry them.

Universal indicator is not used, because it changes colour gradually and has no sharp end-point. The reaction is a **neutralisation**: $\mathrm{NaOH} + \mathrm{HCl} \rightarrow \mathrm{NaCl} + \mathrm{H_2O}$. Ammonium salts are made the same way from aqueous ammonia: $2\mathrm{NH_3} + \mathrm{H_2SO_4} \rightarrow (\mathrm{NH_4})_2\mathrm{SO_4}$.

Insoluble salts: precipitation (Extended)

Extended only. An insoluble salt such as barium sulfate, lead(II) sulfate, silver chloride or lead(II) iodide is made by mixing **two solutions**: one containing its **metal ion** and one containing its **negative ion**. The ions meet and the insoluble salt forms at once as a solid, a **precipitate**.

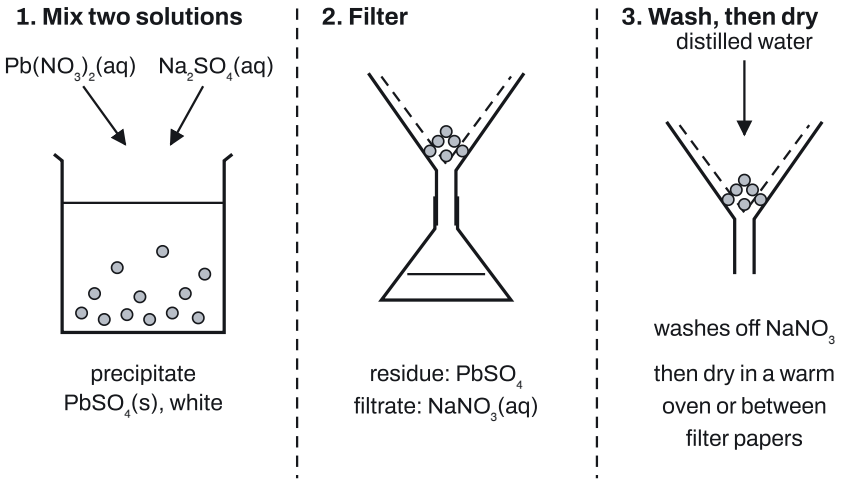
To make lead(II) sulfate you need Pb^{2+} ions (from lead(II) nitrate, because all nitrates dissolve) and $\mathrm{SO_4^{2-}}$ ions (from sodium sulfate, because all sodium salts dissolve, or from dilute sulfuric acid):

$$\mathrm{Pb(NO_3)_2(aq)} + \mathrm{Na_2SO_4(aq)} \rightarrow \mathrm{PbSO_4(s)} + 2\mathrm{NaNO_3(aq)}$$

The **ionic equation** shows only the ions that form the solid:

$$\mathrm{Pb^{2+}(aq) + SO_4^{2-}(aq) \rightarrow PbSO_4(s)}$$

1. **Mix** the two solutions and stir.
2. **Filter.** The precipitate is the **residue**; the soluble by-product (here sodium nitrate) is in the **filtrate**.
3. **Wash** the residue with **distilled water**, to rinse away the soluble by-product still wetting it.
4. **Dry** it in a warm oven or between filter papers.



There is **no crystallisation** step: the salt is already a solid. If the washing step is missed, the soluble by-product dries onto the precipitate, so the "pure" salt weighs more than it should. Useful colours: barium sulfate, lead(II) sulfate, silver chloride and barium carbonate are **white**; silver bromide is **cream**; silver iodide and lead(II) iodide are **yellow**.

You can't make an insoluble salt from an acid and an excess of solid: the salt would be mixed with the leftover solid (and often coats it, stopping the reaction), and filtering would not separate them.

Hydrated and anhydrous salts

Many salts crystallise with water **chemically combined** in the crystal. A **hydrated** substance is chemically combined with water; an **anhydrous** substance contains **no water**. Note the difference from an **aqueous** solution, where the salt is **dissolved** in water.

Extended only: the water in the crystals is the **water of crystallisation**: the water molecules present in hydrated crystals. The formula shows how many per formula unit, after a dot:

- **hydrated copper(II) sulfate,** $\mathrm{CuSO_4 \cdot 5H_2O}$ (blue); heating it gives white anhydrous $\mathrm{CuSO_4}$: $\mathrm{CuSO_4 \cdot 5H_2O \rightarrow CuSO_4 + 5H_2O}$;
- **hydrated cobalt(II) chloride,** $\mathrm{CoCl_2 \cdot 6H_2O}$ (pink); anhydrous $\mathrm{CoCl_2}$ is blue.

Adding water turns the anhydrous forms back (white to blue for copper(II) sulfate, blue to pink for cobalt(II) chloride), which is why they are used to test for water.

Key facts and formulas

"Given" means the exam paper prints it. "Learn it" means you must remember it. Chemistry 0620 has no formula sheet; relative atomic masses come from the Periodic Table printed in the paper.

Formula	Status
Solubility rules: Na, K, $\mathrm{NH_4}$ salts and all nitrates soluble; chlorides soluble except $\mathrm{PbCl_2}$, AgCl; sulfates soluble except $\mathrm{BaSO_4}$, $\mathrm{CaSO_4}$, $\mathrm{PbSO_4}$; carbonates insoluble except Na, K, $\mathrm{NH_4}$; hydroxides insoluble except Na, K, $\mathrm{NH_4}$ ($\mathrm{Ca(OH)_2}$ partly)	Learn it
Soluble salt: acid + excess metal, base or carbonate → filter → heat to the point of crystallisation → cool → filter → dry	Learn it
Na, K, $\mathrm{NH_4}$ salts: titrate acid and alkali with indicator → repeat without indicator → crystallise	Learn it
Extended only: insoluble salt: mix two solutions → filter → wash residue with distilled water → dry	Learn it
Residue = solid left on the filter paper; filtrate = liquid that passes through	Learn it
Saturated solution: maximum concentration of solute dissolved in the solvent at a stated temperature	Learn it
Hydrated = chemically combined with water; anhydrous = contains no water	Learn it
Extended only: water of crystallisation = water molecules present in hydrated crystals; $\mathrm{CuSO_4 \cdot 5H_2O}$, $\mathrm{CoCl_2 \cdot 6H_2O}$	Learn it
moles $= \dfrac{\text{mass}}{M_r}$; moles $= \dfrac{\text{concentration} \times \text{volume in cm}^3}{1000}$	Learn it
Water of crystallisation: $x = \dfrac{\text{moles of water}}{\text{moles of anhydrous salt}}$	Learn it

How to do it

In the questions we have tagged for this topic (111 questions, 2021–2025, counting 10 multiple-choice questions that appear in more than one paper), the types came up this often. 33 of the 111 questions mix types, so the counts add up to 191.

Question type	Questions
Describe or order the steps to make a soluble salt from acid + excess solid	47
Describe making an insoluble salt by precipitation	25
Write the equation for a salt-making reaction	23

Question type	Questions
Use the solubility rules	19
Choose the reactants or method for a named salt	17
Show or explain that the solid is in excess	14
Define hydrated, anhydrous and water of crystallisation	14
Explain saturated solution and why crystals form on cooling	10
Make a soluble salt by titration	8
Find x (water of crystallisation) from heating masses	7
Mole calculations in a salt preparation	7

1. Describe or order the steps to make a soluble salt from acid + excess solid (47 questions)

1. **Add excess** solid to the (warm) dilute acid, stirring, until no more dissolves.
2. **Filter** to remove the excess solid.
3. **Heat the filtrate to the point of crystallisation** (until it is saturated, or until crystals form on a glass rod dipped in it).
4. **Leave to cool** so crystals form.
5. **Filter off the crystals** (or pick them out) and **dry them between filter papers**.

When a question gives prompts ("remove the excess", "crystallise", "dry"), write one clear sentence for each, naming the process **and** saying how it is done.

Variations you will meet:

- **Put the steps in order** (multiple choice): the order is always react → filter → evaporate (partly) → cool and crystallise → dry. Reject options that "evaporate to dryness", that crystallise **before** filtering, or that "wash and dry the residue" (the residue is the unreacted solid, not the salt).
- **Name the residue and the filtrate**: the residue is the **excess solid** (for example magnesium carbonate); the filtrate is the **aqueous salt** (for example aqueous zinc sulfate).
- **Name the separation method**: filtration removes the excess solid; crystallisation gets the salt out of solution.
- **Missing step**: a method that goes straight from filtering to drying has left out the heating to the point of crystallisation and cooling.
- **Full description with an equation** (Paper 4): name the acid and solid, write the balanced equation, then give every step.

2. Describe making an insoluble salt by precipitation (25 questions)

Extended only.

1. **Choose two soluble solutions**: one with the metal ion (use its **nitrate**) and one with the negative ion (use its **sodium or potassium salt**, or the dilute acid).
2. **Mix** them; the salt forms as a precipitate.
3. **Filter**; the salt is the **residue**.
4. **Wash** the residue with **distilled water**.
5. **Dry** it (warm oven or between filter papers).

Variations you will meet:

- **"Which process is not used?"**: **crystallisation**. The salt is already a solid.
- **"Which mixture gives lead(II) sulfate (or chloride) by filtration?"**: two solutions, such as aqueous lead(II) nitrate with dilute sulfuric acid (or dilute hydrochloric acid for lead(II) chloride). A **solid** reactant such as lead(II) carbonate or lead(II) hydroxide is wrong: the product would be mixed with it.
- **"The dry mass is higher than expected"**: the residue was not washed, so the soluble by-product (such as sodium chloride or sodium nitrate) was left in it.
- **Name the method** ("precipitation"), the general term for the solid on the filter paper ("residue"), or the **colour** of the precipitate.
- **Name other reactants**: another soluble salt of the metal (barium nitrate instead of barium chloride), another soluble salt of the negative ion (potassium sulfate instead of sodium sulfate), or another insoluble salt that could be made the same way (barium carbonate, from barium nitrate and sodium carbonate).
- **Recognise precipitation** from an equation: two aqueous reactants giving one product with (s).

3. Write the equation for a salt-making reaction (23 questions)

1. **Pick the pattern**: metal → salt + hydrogen; oxide or hydroxide → salt + water; carbonate → salt + water + carbon dioxide; alkali → salt + water; two solutions → precipitate + soluble salt.
2. **Write the formulae** and **balance**. For example $\mathrm{MgO} + 2\mathrm{HNO_3} \rightarrow \mathrm{Mg(NO_3)_2} + \mathrm{H_2O}$.
3. **State symbols** if asked: solids (s), solutions (aq), water (l), gases (g). The salt made in a soluble-salt preparation is (aq); a precipitate is (s).

Ionic equations for precipitation (Extended): write **only** the two ions that make the solid on the left and **only** the solid on the right, then the state symbols. $\mathrm{Ag^+(aq)} + \mathrm{Cl^-(aq)} \rightarrow \mathrm{AgCl(s)}$; $\mathrm{Pb^{2+}(aq)} + 2\mathrm{I^-(aq)} \rightarrow \mathrm{PbI_2(s)}$.

Variations you will meet:

- **Word equations** (Paper 3): name every product; for a carbonate there are **three** (salt, water, carbon dioxide).
- **Complete an equation** where one product is given: for $\mathrm{BaCl_2} + \mathrm{Na_2SO_4}$ the other product is $2\mathrm{NaCl(aq)}$.
- **Unusual products**: with too little alkali, sulfuric acid can give a hydrogensulfate: $\mathrm{KOH} + \mathrm{H_2SO_4} \rightarrow \mathrm{KHSO_4} + \mathrm{H_2O}$. Follow the formula the question gives.

4. Use the solubility rules (19 questions)

1. **Split each salt** into its metal (or ammonium) and its negative ion.
2. **Check Na, K, $\mathrm{NH_4}$ and nitrate first**: if either part is one of these, the salt is soluble.
3. **Otherwise use the row for the negative ion** and its exceptions.

Variations you will meet:

- **"Which three compounds are all insoluble?"**: test every compound; one soluble one makes the option wrong. Barium carbonate, lead(II) sulfate and zinc hydroxide are all insoluble; potassium sulfate and ammonium carbonate are not.
- **"Which pair of solutions gives a precipitate?"**: first check that both reactants are soluble (a solid such as barium sulfate can't be one "solution"), then swap partners and see whether a new pair is insoluble. $\mathrm{Ba(NO_3)_2 + K_2CO_3}$ gives barium carbonate, a precipitate; $\mathrm{KCl + Zn(NO_3)_2}$ gives nothing.
- **A table of results** (which salts dissolve when stirred with water): match each result to a salt, for example calcium sulfate and lead(II) chloride stay as solids.
- **"Name another insoluble chloride"**: the only two are lead(II) chloride and silver chloride.
- **Given rules**: some questions print their own rules; use those.

5. Choose the reactants or method for a named salt (17 questions)

1. **Insoluble salt?** → precipitation from two soluble salts.
2. **Sodium, potassium or ammonium salt?** → titration of the acid with the alkali (or aqueous ammonia).
3. **Any other soluble salt?** → the acid + an excess of the metal's oxide, hydroxide or carbonate (or the metal, if it is magnesium, zinc or iron).
4. **The acid** gives the negative ion: sulfuric → sulfate, hydrochloric → chloride, nitric → nitrate.

Variations you will meet:

- **"Name another substance that makes the same salt"** with the acid: if the question used the carbonate, give the **oxide** or **hydroxide** (for copper(II) sulfate: copper(II) oxide or copper(II) hydroxide). Give exactly the number of answers asked for.
- **"Which methods make copper(II) chloride?"**: acid + copper(II) oxide or carbonate, **not** copper + acid (no reaction).
- **Titration reagent** for a sodium salt: the **alkali** (sodium hydroxide), not an insoluble solid.
- **"Pure crystals do not form if calcium is used instead of zinc"**: calcium also reacts with the water, making calcium hydroxide, so the excess metal doesn't stay as a solid you can filter off.
- **Which reactant pair makes calcium sulfate** (insoluble): two solutions, for example calcium nitrate and sodium sulfate, not calcium carbonate + sulfuric acid.

6. Show or explain that the solid is in excess (14 questions)

1. **Observations that show excess**: some **solid remains** (it stops disappearing); with a metal or carbonate, **no more fizzing** when more is added.
2. **Why excess?** To make sure **all the acid reacts** (the acid is the limiting reactant), so no acid is left in the salt; the unreacted solid is removed by filtering.

Variations you will meet:

- **"Name the limiting reactant"**: the **acid**.
- **"Give two observations in step 1"** while it reacts: fizzing (bubbles), the solid disappears, and the solution changes colour (**blue** for copper salts).
- **"Why does the reaction stop?"**: all the acid has been used up.

7. Define hydrated, anhydrous and water of crystallisation (14 questions)

1. **Hydrated**: chemically combined with water. **Anhydrous**: contains no water.
2. **Extended only: water of crystallisation**: the water molecules present in hydrated crystals.
3. **Know the formulae**: $\mathrm{CuSO_4 \cdot 5H_2O}$ and $\mathrm{CoCl_2 \cdot 6H_2O}$.

Variations you will meet:

- **Name the term** from a description ("a substance chemically combined with water" → hydrated).
- **Match forms to descriptions**: aqueous = dissolved in water; hydrated solid = chemically combined with water; anhydrous solid = containing no water.
- **Pick x and y** in $\mathrm{CuSO_4 \cdot} x\mathrm{H_2O}$ and $\mathrm{CoCl_2 \cdot} y\mathrm{H_2O}$: 5 and 6.
- **What is left after heating to constant mass**: the anhydrous salt, for example $\mathrm{ZnSO_4}$.
- **Colours**: hydrated copper(II) sulfate blue, anhydrous white; hydrated cobalt(II) chloride pink, anhydrous blue.

8. Explain saturated solution and why crystals form on cooling (10 questions)

1. **Saturated solution**: a solution containing the **maximum concentration of solute** dissolved in the solvent **at a stated temperature**. Both parts earn a mark.
2. **Why crystals form on cooling**: **solubility decreases as temperature decreases**, so the solution can't hold as much salt.

Variations you will meet:

- **"State the term"** for a solution holding the most salt that will dissolve at that temperature: **saturated**.
- **"How do you know it is saturated?"**: crystals form on a glass rod dipped in and taken out (or on a drop placed on a cold surface).
- **"Effect on the mass of crystals if the solution is left to dry without heating"**: **none**; it just takes longer.
- **Saturated limewater**: stir excess calcium hydroxide with water, then filter.

9. Make a soluble salt by titration (8 questions)

1. **Pipette** the alkali into a conical flask; add a few drops of **methyl orange** or **thymolphthalein**.
2. **Add the acid from a burette** until the end-point (methyl orange yellow → orange; thymolphthalein blue → colourless). Note the volume.
3. **Repeat without the indicator**, using the same volumes.
4. **Heat to the point of crystallisation, cool, filter and dry** the crystals.

Variations you will meet:

- **"Name the type of reaction"**: neutralisation (it is exothermic).
- **"Name the apparatus used to add the acid"**: a burette.
- **"Why not universal indicator?"**: it has too many colour changes (it changes gradually), so there is no sharp end-point.
- **Acid in the flask instead**: you can also pipette the acid into the flask and add the alkali from the burette. Then the colours run the other way: thymolphthalein goes from **colourless to blue**, methyl orange from **red to orange** (yellow if you overshoot).
- **"The crystals are coloured, not white"**: the indicator was left in; repeat without it.
- **Choosing the other reagent** for a sodium or potassium salt: the alkali (sodium hydroxide or potassium hydroxide).

10. Find x (water of crystallisation) from heating masses (7 questions)

1. **Moles of anhydrous salt** $= \dfrac{\text{mass left}}{M_r}$.
2. **Mass of water** $=$ mass of hydrated salt $-$ mass of anhydrous salt.
3. **Moles of water** $= \dfrac{\text{mass of water}}{18}$.
4. $x = \dfrac{\text{moles of water}}{\text{moles of anhydrous salt}}$, a whole number.

Variations you will meet:

- **"How do you know all the water has gone?"**: heat, cool and weigh, then **repeat until the mass stays the same** (heating to constant mass).
- **Moles of water given**: skip steps 2 and 3 and divide straight away.
- **Name the crystals** that contain water of crystallisation: hydrated.

11. Mole calculations in a salt preparation (7 questions)

These use the methods of The mole. The usual steps:

1. **Moles from a mass**: $n = \dfrac{m}{M_r}$; **mass from moles**: $m = n \times M_r$. For example 0.0200 mol of $\mathrm{CuSO_4{\cdot}5H_2O}$ ($M_r = 250$) is 5.00 g.

2. **Moles in a solution**: $n = \dfrac{c \times V}{1000}$ with V in cm^3; **volume** $V = \dfrac{n}{c} \times 1000$; **concentration** $c = \dfrac{n}{V} \times 1000$.
3. **Use the equation's ratio**: $2\text{NaOH} + \text{H}_2\text{SO}_4 \rightarrow \text{Na}_2\text{SO}_4 + 2\text{H}_2\text{O}$, so 0.0100 mol of NaOH needs 0.00500 mol of acid.
4. **Expected (maximum) mass of product** = moles of the limiting reactant × ratio × M_r of the product.
5. **g/dm³** = mol/dm³ $\times M_r$.

Worked examples

The bold codes show where marks are earned, as in Cambridge mark schemes. The number is how many marks.

- **M1** method mark: for a correct method, even if the numbers slip.
- **A1** accuracy mark: for a correct answer; it needs the M mark before it.
- **B1** mark for a correct result on its own, with no method needed.

Example 1

Nickel(II) hydroxide, Ni(OH)_2, is a green solid that does not dissolve in water. A student wants crystals of nickel(II) nitrate and writes this method.

- **1** Add 2 g of nickel(II) hydroxide to $50\ \text{cm}^3$ of dilute nitric acid and stir until all the solid has dissolved.
- **2** Pour the solution into an evaporating basin and heat it until no liquid is left.
- **3** Scrape out the solid and put it in a bottle.

(a) Complete the word equation: nickel(II) hydroxide + nitric acid → ... + ... **[2]**

(b) Step 1 and step 2 each contain a mistake. For each step, say what is wrong and what the student should do instead. **[4]**

(c) The student says: "I will know the solid is in excess when the fizzing stops." Explain why this won't work here, and give the observation the student should look for instead. **[2]**

(d) Name **one** other nickel compound that could be reacted with dilute nitric acid to make the same salt. **[1]**

(a) Nickel(II) nitrate **B1** + water. **B1**

(b) Step 1: if all the solid dissolves, some acid may be left over, and it would end up in the crystals. **B1** Keep adding nickel(II) hydroxide until some is left undissolved (in excess), then filter off the excess. **B1** Step 2: heating until no liquid is left drives off the water of crystallisation and leaves any soluble impurities in the solid. **B1** Heat only to the point of crystallisation (until crystals form on a glass rod dipped in), leave to cool, then filter off the crystals and dry them between filter papers. **B1**

(c) A hydroxide reacts with an acid to give only a salt and water: no gas is made, so there is no fizzing at any stage. **B1** Look for green solid that stays in the flask and no longer disappears when stirred. **B1**

(d) Nickel(II) oxide (or nickel(II) carbonate). **B1**

Example 2

Extended. A student makes calcium carbonate, $\mathrm{CaCO_3}$, by mixing $30.0\ \mathrm{cm^3}$ of $0.500\ \mathrm{mol/dm^3}$ aqueous calcium chloride with $20.0\ \mathrm{cm^3}$ of $1.00\ \mathrm{mol/dm^3}$ aqueous potassium carbonate. (M_r: $\mathrm{CaCO_3}$ = 100)

(a) Use the solubility rules to explain why calcium carbonate forms as a solid, while the other product stays in solution. **[2]**

(b) Write the ionic equation for the reaction. Include state symbols. **[3]**

(c) Show which reactant is in excess. **[2]**

(d) Calculate the maximum mass of calcium carbonate that can form. **[1]**

(e) Name **three** ions left in the filtrate. **[2]**

(f) Describe how the student gets a pure, dry sample of calcium carbonate from the mixture. **[2]**

(a) Carbonates are insoluble except those of sodium, potassium and ammonium, so calcium carbonate precipitates. **B1** The other product, potassium chloride, is soluble (all potassium salts are). **B1**

(b) $\mathrm{Ca^{2+}(aq) + CO_3^{2-}(aq) \rightarrow CaCO_3(s)}$. **B1** for $\mathrm{CaCO_3}$ as the only product, **B1** for only $\mathrm{Ca^{2+}}$ and $\mathrm{CO_3^{2-}}$ on the left, **B1** for the state symbols.

(c) Moles of $\mathrm{CaCl_2} = \dfrac{0.500 \times 30.0}{1000} = 0.0150$; moles of $\mathrm{K_2CO_3} = \dfrac{1.00 \times 20.0}{1000} = 0.0200$. **M1** They react $1:1$, so the potassium carbonate is in excess (0.0050 mol left over). **A1**

(d) The calcium chloride is used up, so $0.0150 \times 100 = 1.50\ \mathrm{g}$. **B1**

(e) $\mathrm{K^+}$, $\mathrm{Cl^-}$ **B1** and the excess $\mathrm{CO_3^{2-}}$. **B1**

(f) Filter, and wash the residue with distilled water to rinse off the soluble salts. **B1** Dry it in a warm oven or between filter papers. **B1**

Example 3

Extended. A student wants to make $2.55\ \mathrm{g}$ of sodium nitrate, $\mathrm{NaNO_3}$ (M_r = 85), by titration. $25.0\ \mathrm{cm^3}$ of $1.20\ \mathrm{mol/dm^3}$ nitric acid is pipetted into a conical flask with a few drops of thymolphthalein. Aqueous sodium hydroxide of concentration $0.800\ \mathrm{mol/dm^3}$ is added from a burette.

(a) Write the symbol equation for the reaction. Include state symbols. **[2]**

(b) Show that the nitric acid can make at most $2.55\ \mathrm{g}$ of sodium nitrate. **[2]**

(c) Calculate the volume of sodium hydroxide needed to reach the end-point. **[2]**

(d) State the colour of the mixture in the flask at the start and at the end-point. **[2]**

(e) A friend suggests adding an excess of solid sodium carbonate to the acid and filtering. Explain why this gives an impure product. **[1]**

(f) Describe how the student makes pure, dry sodium nitrate crystals once the volume in (c) is known. **[3]**

(a) $\mathrm{HNO_3(aq) + NaOH(aq) \rightarrow NaNO_3(aq) + H_2O(l)}$ **B1** state symbols **B1**

(b) Moles of $\mathrm{HNO_3} = \dfrac{1.20 \times 25.0}{1000} = 0.0300$. **M1** The ratio is $1:1$, so $0.0300 \times 85 = 2.55$ g. **A1**

(c) 0.0300 mol of NaOH is needed. **M1** $V = \dfrac{0.0300}{0.800} \times 1000 = 37.5\ \mathrm{cm^3}$. **A1**

(d) Colourless at the start (acid) **B1**; blue at the end-point. **B1**

(e) Sodium carbonate is soluble, so the excess dissolves too and can't be filtered off; it would crystallise with the sodium nitrate. **B1**

(f) Mix $25.0\ \mathrm{cm^3}$ of the acid with $37.5\ \mathrm{cm^3}$ of the alkali again, **without the indicator**. **B1** Heat the solution to the point of crystallisation and leave it to cool. **B1** Filter off the crystals and dry them between filter papers. **B1**

Example 4

Extended. Hydrated strontium chloride is $\mathrm{SrCl_2 \cdot}x\mathrm{H_2O}$. A student heats a sample in a crucible, lets it cool and weighs it, several times. (M_r: $\mathrm{SrCl_2}$ = 159, $\mathrm{H_2O}$ = 18)

	mass of solid / g
before heating	5.34
after heating 1	4.02
after heating 2	3.36
after heating 3	3.18
after heating 4	3.18

(a) Which mass should be used for the anhydrous salt? Explain your choice. **[2]**

(b) Determine the value of x. **[4]**

(c) Another student stops after heating 2 and uses 3.36 g as the mass of anhydrous salt. Would their value of x be too high or too low? Explain. **[2]**

(a) 3.18 g **B1**: the mass stayed the same after two heatings (constant mass), so all the water of crystallisation had gone. **B1**

(b) Moles of $\mathrm{SrCl_2} = \dfrac{3.18}{159} = 0.0200$. **B1**

Mass of water $= 5.34 - 3.18 = 2.16$ g. **B1**

Moles of water $= \dfrac{2.16}{18} = 0.120$. **B1**

$x = \dfrac{0.120}{0.0200} = 6$. **B1** (The crystals are $\mathrm{SrCl_2 \cdot 6H_2O}$.)

(c) Too low. **B1** Some water was still in the solid, so the mass of water lost looks smaller and the moles of "salt" look bigger (here x would come out as about 5.2). **B1**

Common mistakes

- **Heating to dryness.** Examiners see "evaporate all the water" or "heat until dry" every year. Heat only **until the solution is saturated** (point of crystallisation), then let it cool.
- **Writing "crystallise the solution" without saying how.** Describe it: heat until crystals form on a glass rod, then leave to cool.
- **Drying too harshly or vaguely.** "Leave to dry" is too vague, and a hot oven can drive off water of crystallisation. Say "dry between filter papers" (or "in a warm place").
- **Mixing up residue and filtrate.** The residue stays **on** the filter paper; the filtrate goes **through**. In a soluble-salt preparation, you **don't** want the residue: "dry the residue" is wrong.
- **Explaining crystals on cooling by evaporation.** Crystals form because **solubility decreases as the temperature falls**.
- **Leaving half off the saturated-solution definition.** "At a stated temperature" is needed as well as "maximum amount of solute dissolved"; don't swap solute and solvent.
- **Vague "excess" observations.** "Add excess" is not an observation. Say what you **see**: solid remains; no more fizzing.
- **Not using water to wash a precipitate.** "Wash the residue" needs "with (distilled) water".
- **Vague definition of hydrated.** "Contains water", "dissolved in water" or "absorbs water" are not enough: hydrated means **chemically combined** with water.
- **Solubility rules not learned.** Many candidates can't pick out the insoluble compound. The exceptions (lead and silver chlorides; barium, calcium and lead sulfates) are what questions test.
- **Wrong formulae in precipitation equations.** Use the formula given in the question ($\mathrm{PbI_2}$, not PbI), and write sodium nitrate as $\mathrm{NaNO_3}$, without brackets.
- **Giving too many answers.** When asked for **two** compounds, give two: a wrong third one can cancel a right one.

How to get the marks

- **"Describe how to prepare"**: one mark per key step, each with its reason or detail: filter (to remove the excess solid); heat to the point of crystallisation; cool; filter off crystals; dry between filter papers. For precipitation: mix, filter, **wash with distilled water**, dry.
- **"Name"**: give a full chemical name, such as "copper(II) oxide", not just "a base" or "copper".

- **"State"** (a term such as residue, saturated, hydrated, precipitation): just the word, spelled correctly.
- **"Give observations"**: what you would **see** (fizzing, solid disappears, blue solution), not "carbon dioxide is made".
- **Ionic precipitation equations**: marks come separately for the product, the ions on the left and the state symbols, so get each right.
- **Colour changes**: "from ... to ...", both colours, in the right order.
- **Calculations**: show each step the question lists (moles, mass of water, moles of water, x); a wrong earlier answer used correctly still scores later marks. Give answers to 3 significant figures unless told otherwise, and x as a whole number.
- **"Suggest"**: apply the method to a salt you haven't met, using the flowchart and the solubility rules.
- **Multiple choice**: check every part of an option. A method that is right except for "crystallise the residue" or "evaporate to dryness" is wrong.

Quick check

1. Which of these are insoluble in water: zinc carbonate, ammonium sulfate, lead(II) nitrate, barium sulfate, sodium hydroxide, iron(II) hydroxide?
2. Choose a method (titration, excess solid, or precipitation) and the reactants for each salt: (a) iron(II) chloride; (b) ammonium nitrate; (c) calcium sulfate; (d) zinc nitrate.
3. **Extended.** 2.86 g of washing soda crystals, $\mathrm{Na_2CO_3{\cdot}10H_2O}$ ($M_r = 286$), are heated to constant mass. What mass of anhydrous sodium carbonate ($M_r = 106$) is left, and what mass of water is lost?
4. **Extended.** Silver iodide is made from aqueous silver nitrate and aqueous potassium iodide. Write the ionic equation with state symbols, give the colour of the precipitate, and name the soluble product that must be washed off it.
5. A student heats zinc nitrate solution in a basin, then stands the basin in iced water to collect crystals. (a) How can the student tell when to stop heating? (b) Explain why standing the basin in iced water gives a bigger mass of crystals than leaving it to cool on the bench.

Answers

1. Zinc carbonate, barium sulfate and iron(II) hydroxide. (Ammonium and sodium compounds and all nitrates are soluble.)
2. (a) Excess solid: iron, iron(II) oxide or iron(II) carbonate + hydrochloric acid. (b) Titration: aqueous ammonia + nitric acid. (c) Precipitation: for example calcium nitrate solution + sodium sulfate solution. (d) Excess solid: zinc, zinc oxide or zinc carbonate + nitric acid.
3. Moles $= \dfrac{2.86}{286} = 0.0100$, so $0.0100 \times 106 = 1.06$ g of $\mathrm{Na_2CO_3}$ is left, and $2.86 - 1.06 = 1.80$ g of water is lost.
4. $\mathrm{Ag^+(aq) + I^-(aq) \rightarrow AgI(s)}$. The precipitate is yellow. The soluble product is potassium nitrate.
5. (a) Stop when the solution is saturated: dip in a glass rod and take it out; if crystals form on it, stop heating. (b) The solubility of zinc nitrate decreases as the temperature decreases, so at the lower temperature of iced water less salt can stay dissolved and more comes out as crystals.

Past questions to try

- **0620/33/M/J/25 Q5**: part (b), describing how to get pure, dry crystals after reacting an excess of a metal oxide.
- **0620/43/M/J/24 Q5**: barium sulfate by precipitation, then finding x in a hydrated salt.
- **0620/42/F/M/21 Q4**: zinc chloride from excess zinc, saturated solutions, and a titration calculation.
- **0620/43/O/N/21 Q5**: parts (a) and (b), making sodium sulfate crystals by titration.