Chapter: 8: Acids, bases and salts
Book/PDF: Chapter-08.pdf
Pages: 116–133
Exam level: Cambridge IGCSE (0610)

1) Big-picture overview (100–150 words)

This chapter introduces the fundamental concepts of acids, bases, and salts, which are crucial categories of chemical compounds. You'll learn how to identify acids and alkalis using indicators and the pH scale (p. 117), and understand the key difference between a strong acid (fully dissociates) and a weak acid (partially dissociates) (p. 118-119). The chapter focuses heavily on the reactions between acids and bases, known as neutralisation, which produce a salt and water (p. 120). It provides detailed, practical methods for preparing both soluble and insoluble salts in the lab, such as titration and precipitation (p. 121-124). Finally, you will explore hydrated salts, which contain 'water of crystallisation' within their crystal structure, and learn how to test for the presence of common ions (p. 124-126).

2) Syllabus mapping

Outcome code	Outcome description	Where covered (page)
(Not stated)	Define acids as proton donors and bases as proton acceptors (Brønsted-Lowry theory).	p. 118
(Not stated)	Describe the characteristic properties of acids and bases.	p. 116, 119
(Not stated)	Explain the difference between strong and weak acids/bases in terms of dissociation.	p. 118, 119
(Not stated)	Explain the difference between concentrated and dilute solutions.	p. 119
(Not stated)	Describe the use of indicators (litmus, methyl orange, thymolphthalein, universal indicator) and the pH scale.	p. 117
(Not stated)	Define neutralisation and write ionic equations for it.	p. 120, 123
(Not stated)	Describe the preparation of soluble salts using an acid with a metal, insoluble base, or carbonate.	p. 121–123
(Not stated)	Describe the preparation of soluble salts of reactive metals using titration.	p. 122
(Not stated)	Describe the preparation of insoluble salts by precipitation.	p. 124

Outcome code	Outcome description	Where covered (page)
(Not stated)	Know the general solubility rules for common salts.	p. 120
(Not stated)	Describe tests to identify anions: carbonates, chlorides, bromides, iodides, nitrates, and sulfates.	p. 124–125
(Not stated)	Define and explain 'water of crystallisation' in hydrated salts.	p. 126

3) Key terms and definitions

Term	One-sentence definition	First appears (page)	Example/application
Indicator	A substance that changes colour when added to acidic or alkaline solutions (p. 117).	p. 117	Litmus paper turns red in acid; Universal indicator shows a range of colours corresponding to pH (p. 117).
pH scale	A scale from below 0 to 14 that measures the acidity or alkalinity of a substance (p. 117).	p. 117	pH < 7 is acidic, pH = 7 is neutral, pH > 7 is alkaline (p. 117).
Acid	A substance that is a proton (H^+ ion) donor (p. 118).	p. 118	Hydrochloric acid (HCl) donates an H^+ ion when dissolved in water (p. 118).
Base	A substance that is a proton (H^+ ion) acceptor (p. 118).	p. 118	Sodium hydroxide ($NaOH$) accepts an H^+ ion during neutralisation (p. 120).
Strong acid	An acid that is completely dissociated (ionised) in aqueous solution (p. 119).	p. 119	Hydrochloric acid (HCl), sulfuric acid (H_2SO_4), nitric acid (HNO_3) (p. 119).
Weak acid	An acid that is partially dissociated in aqueous solution (p. 119).	p. 119	Ethanoic acid (CH_3COOH), citric acid, carbonic acid (p. 119).
Neutralisation	A reaction between an acid and a base to produce a salt and water (p. 120).	p. 120	The core reaction is $H^+(aq) + OH^-(aq) rightarrow H\_2O(l)$ (p. 120).

Term	One-sentence definition	First appears (page)	Example/application
Salt	A compound formed when the hydrogen ions of an acid are replaced by metal ions or the ammonium ion (p. 120).	p. 120	Sodium chloride ($NaCl$), copper(II) sulfate ($CuSO_4$) (p. 120).
Alkali	A base that is soluble in water and produces hydroxide ions (OH^-) in solution (p. 122).	p. 122	Sodium hydroxide ($NaOH$), potassium hydroxide (KOH) (p. 122).
Titration	A technique used to find the exact volume of one solution needed to react completely with a measured volume of another solution (p. 122, 128).	p. 122	Used to prepare salts of reactive metals (like sodium or potassium) or to find an unknown concentration (p. 122, 128).
Precipitation	A reaction where two soluble salts are mixed to form an insoluble salt (the precipitate) (p. 124).	p. 124	Mixing barium chloride and sodium sulfate solutions to make solid barium sulfate (p. 124).
Hydrate	A salt which incorporates water into its crystal structure (p. 126).	p. 126	Copper(II) sulfate pentahydrate ($CuSO_4.5H_2O$) is a blue crystal (p. 126).
Anhydrous	A substance containing no water (p. 126).	p. 126	Anhydrous copper(II) sulfate ($CuSO_4$) is a white powder (p. 126).
Water of crystallisation	The water molecules present within the crystal structure of a hydrated salt (p. 126).	p. 126	The '$5H_2O$' in $CuSO_4.5H_2O$ (p. 126).

4) Core concepts explained

8.1 Acids and Alkalis (p. 116)

- **Acids** are substances with a pH less than 7 (p. 117). They taste sour (like lemons) but should never be tasted in a lab (p. 117).
- **Alkalis** are soluble bases with a pH greater than 7 (p. 117, 122). Many cleaning products are alkaline (p. 117).

- **Indicators** are dyes that change colour to show if a substance is acidic or alkaline (p. 117). Universal indicator shows a spectrum of colours to estimate the pH value (p. 117).
- **The pH scale** runs from <0 to 14, with 7 being neutral (like pure water) (p. 117). A pH meter provides a precise digital reading (p. 118).
- **The Brønsted-Lowry theory** defines an acid as a proton (H^+) donor and a base as a proton (H^+) acceptor (p. 118). For an acid to act as a donor, a base must be present to accept the proton (p. 118).

Feature	Strong Acid	Weak Acid	Exam Note
Dissociation in water	Complete/full dissociation (p. 118).	Partial/incomplete dissociation (p. 119).	The reversible arrow ($rightleftharpoons$) is used for weak acids (p. 119).
H^+ **concentration**	High concentration of H^+ ions (p. 119).	Low concentration of H^+ ions (p. 119).	For the same molar concentration.
pH value	Low pH (e.g., 0-2) (p. 119).	Higher pH (e.g., 3-6) (p. 119).	Still less than 7.
Conductivity	Good electrical conductor (p. 119).	Poor electrical conductor (p. 119).	Due to the number of mobile ions.
Rate of reaction	Fast reaction with metals, bases, etc. (p. 119).	Slow reaction with metals, bases, etc. (p. 119).	Rate depends on H^+ concentration.
Examples	HCl, H_2SO_4, HNO_3 (p. 119).	CH_3COOH (ethanoic acid), Citric acid (p. 119).	Be able to name one of each.

- **Concentration vs. Strength:** These are different concepts (p. 119).
  - **Strength** refers to the degree of dissociation of the acid (p. 119). A strong acid is always strong, even when dilute (p. 119).
  - **Concentration** refers to the amount of acid dissolved in a certain volume of water (p. 119). You can have a concentrated weak acid or a dilute strong acid.

8.2 Formation of salts (p. 120)

- A **salt** is formed when an acid reacts with a base in a neutralisation reaction (p. 120).
- The name of the salt has two parts: the first part comes from the metal in the base (e.g., 'sodium' from sodium hydroxide), and the second part comes from the acid (p. 121).
  - Hydrochloric acid → **Chlorides**
  - Sulfuric acid → **Sulfates**
  - Nitric acid → **Nitrates**
  - Ethanoic acid → **Ethanoates**
- **Solubility rules** are essential for choosing the correct preparation method (p. 120):

- **Always Soluble:** All sodium, potassium, and ammonium salts; all nitrates (p. 120).
- **Mostly Soluble:** Chlorides (except lead and silver); Sulfates (except barium, calcium, lead) (p. 120).
- **Mostly Insoluble:** Carbonates and hydroxides (except sodium, potassium, ammonium) (p. 120).

8.3 & 8.4 Methods of Preparing Salts (p. 121-124)

Method	Reactants	Products	Salt Type	Key Steps
Acid + Metal	Acid + Excess Metal	Salt + Hydrogen	Soluble	1. Add excess metal to acid. 2. Filter off unreacted metal. 3. Evaporate water to crystallise. (p. 121)
Acid + Insoluble Base	Acid + Excess Insoluble Base (oxide/hydroxide)	Salt + Water	Soluble	1. Add excess base to warm acid. 2. Filter off unreacted base. 3. Evaporate to crystallise. (p. 123)
Acid + Insoluble Carbonate	Acid + Excess Insoluble Carbonate	Salt + Water + Carbon Dioxide	Soluble	1. Add excess carbonate to acid. 2. Filter off unreacted carbonate. 3. Evaporate to crystallise. (p. 122)
Acid + Alkali (Titration)	Acid + Alkali (soluble base)	Salt + Water	Soluble	1. Titrate with indicator to find end-point. 2. Repeat without indicator. 3. Evaporate to crystallise. (p. 122)
Precipitation	Soluble Salt A + Soluble Salt B	Insoluble Salt + Soluble Salt	Insoluble	1. Mix two soluble salt solutions. 2. Filter the precipitate. 3. Wash with distilled water and dry. (p. 124)

8.6 Water of Crystallisation (p. 126)

- Many salts crystallise from a solution with water molecules chemically bonded into the crystal lattice. This is called a **hydrate** (p. 126).
- The water is called **water of crystallisation** (p. 126). Its formula is shown with a dot, e.g., $CuSO_4 \cdot 5H_2O$.
- Heating a hydrate drives off the water, leaving the **anhydrous** salt (p. 126). This is often accompanied by a colour change.
- **Example:** Blue hydrated copper(II) sulfate ($CuSO_4 \cdot 5H_2O$) turns into white anhydrous copper(II) sulfate ($CuSO_4$) upon heating (p. 126).

- The reverse reaction can be used as a chemical test for water: adding water to anhydrous copper(II) sulfate turns it blue (p. 127).

5) Diagrams and micrographs (figures)

Figure	Description	Labels / Key Components
Fig 8.3 (p. 117)	Shows various chemical indicators in beakers, displaying different colours in acidic and alkaline solutions.	Beakers, dropper bottles, coloured solutions (e.g., pink, yellow). Illustrates how indicators work.
Fig 8.4 (p. 118)	The pH scale from 0 to 14, with colours from universal indicator and examples of common substances at their respective pH values.	Strong acid (red), Neutral (green), Strong alkali (purple). Examples: Lemon juice (pH 2), Water (pH 7), Bleach (pH 13).
Fig 8.5 (p. 118)	A digital pH meter with its probe in a beaker of solution, giving a precise numerical reading.	Digital display, pH electrode (probe), beaker.
Fig 8.7 (p. 121)	Filtration apparatus used to separate a solid from a liquid.	Funnel, filter paper, beaker/crucible to collect the filtrate, glass rod to guide the liquid.
Fig 8.8 (p. 122)	Evaporation setup to concentrate a solution.	Evaporating dish on a gauze over a tripod, heated by a Bunsen burner.
Fig 8.9 (p. 123)	A titration in progress, showing acid being added from a burette to an alkali in a flask.	Burette (held by a stand), conical flask. Hand swirling the flask.
(p. 127)	Apparatus for a titration.	**Pipette with safety filler:** for accurately measuring a volume of alkali. **Burette:** for adding a variable, but accurately known, volume of acid. **Conical flask:** holds the alkali and allows swirling. **Stand:** holds the burette vertically.
Fig 8.12 (p. 124)	Shows precipitation. A clear solution being poured into another, forming a white solid (precipitate).	Test tube, beaker/bottle. Shows the formation of an insoluble salt.

Figure	Description	Labels / Key Components
Fig 8.16 (p. 127)	Test for water. A white powder (anhydrous copper(II) sulfate) at the bottom of a test tube turning blue where water has been added.	Test tube. Demonstrates the reversible reaction of a hydrated salt.

6) Processes and cycles

Preparation of a Soluble Salt (e.g., Copper(II) Sulfate) (p. 123)

This uses the "acid + insoluble base" method.

1. **Reaction:** Gently warm some dilute sulfuric acid in a beaker. Add excess copper(II) oxide (a black powder) bit by bit, stirring until no more dissolves. This ensures all the acid has reacted (p. 123).
2. **Filtration:** Filter the hot mixture to remove the unreacted excess copper(II) oxide. The blue solution collected is copper(II) sulfate (p. 123, similar to Fig 8.7).
3. **Evaporation:** Gently heat the solution in an evaporating dish to boil off about half the water, concentrating the solution (p. 123, similar to Fig 8.8). Test for saturation by dipping a cold glass rod in; crystals should form on it (p. 123).
4. **Crystallisation:** Leave the saturated solution to cool slowly. As it cools, the solubility decreases and blue crystals of hydrated copper(II) sulfate will form (p. 123).
5. **Drying:** Filter the crystals from the remaining solution and dry them between sheets of filter paper (p. 123).

- **Inputs:** Dilute sulfuric acid (H_2SO_4), Copper(II) oxide (CuO).
- **Outputs:** Copper(II) sulfate ($CuSO\_4$), Water (H_2O).
- **Word Equation:** Sulfuric acid + Copper(II) oxide → Copper(II) sulfate + Water (p. 123).
- **Chemical Equation:** $H\_2SO\_4(aq) + CuO(s) rightarrow CuSO\_4(aq) + H\_2O(l)$ (p. 123).

Preparation of an Insoluble Salt (e.g., Barium Sulfate) (p. 124)

This uses the precipitation method.

1. **Choose Reactants:** Select two soluble salts. One must contain the cation (positive ion) of the desired salt (e.g., Barium from Barium chloride), and the other must contain the anion (negative ion) (e.g., Sulfate from Sodium sulfate) (p. 124).
2. **Mixing:** Mix the two solutions together in a beaker (p. 124). A solid white precipitate of barium sulfate will form instantly.
3. **Filtration:** Filter the mixture to separate the solid precipitate from the solution.
4. **Washing & Drying:** Wash the precipitate on the filter paper with distilled water to remove any impurities from the remaining soluble salt solution. Then, dry the precipitate in a warm oven or between filter paper.

- **Inputs:** Barium chloride solution ($BaCl\_2$), Sodium sulfate solution (Na_2SO_4).
- **Outputs:** Barium sulfate (solid precipitate, $BaSO\_4$), Sodium chloride solution ($NaCl$).
- **Chemical Equation:** $BaCl\_2(aq) + Na\_2SO\_4(aq) rightarrow BaSO\_4(s) + 2NaCl(aq)$ (p. 124).
- **Ionic Equation:** $Ba^{2+}(aq) + SO\_4^{2-}(aq) rightarrow BaSO\_4(s)$ (p. 124).

Qualitative Analysis: Testing for Anions (p. 124-125)

Anion	Test Procedure	Positive Result
Carbonate (CO_3^{2-})	Add a few drops of any dilute acid (e.g., HCl). Bubble any gas produced through limewater.	Effervescence (fizzing). Gas turns limewater milky/cloudy (p. 125).
Sulfate (SO_4^{2-})	Add a few drops of dilute hydrochloric acid, followed by a few drops of barium chloride solution.	A white precipitate (barium sulfate) forms (p. 124).
Chloride (Cl^-)	Add a few drops of dilute nitric acid, followed by a few drops of silver nitrate solution.	A white precipitate (silver chloride) forms (p. 125).
Bromide (Br^-)	Add dilute nitric acid, then silver nitrate solution.	A cream precipitate (silver bromide) forms (p. 125).
Iodide (I^-)	Add dilute nitric acid, then silver nitrate solution.	A yellow precipitate (silver iodide) forms (p. 125).
Nitrate (NO_3^-)	Add aqueous sodium hydroxide, then a small piece of aluminium foil, and warm gently.	Pungent gas (ammonia) is produced, which turns damp red litmus paper blue (p. 125).

7) Formulae and calculations

Quantity	Formula	Units	Typical values	Worked example (from p. 130)
Moles in solution	$moles = fracvolume1000 timesconcentration$ (p. 129)	mol	10^{-2} to 10^{-4} mol	Moles of $0.10\tilde{m}ol/dm^3$ HCl in $21.0\tilde{c}m^3$: $moles = frac21.01000 times0.10 = 0.0021\tilde{m}ol$ (p. 129)

Quantity	Formula	Units	Typical values	Worked example (from p. 130)
Titration calculation	$fracM\_1V\_1n\_1 = fracM\_2V\_2n\_2$ *(M is concentration, V is volume, n is the mole ratio from the balanced equation. The book uses 'M' for the mole ratio instead of 'n')* (p. 130)	Concentration: mol/dm^3 Volume: cm^3	Acid/alkali concentrations are often 0.1-1.0 mol/dm^3. Volumes are typically ~25 cm^3.	**Problem:** Find the concentration of $25.0\tilde{c}m^3$ of $NaOH$ that reacts with $21.0\tilde{c}m^3$ of $0.10\tilde{m}ol/dm^3\tilde{H}Cl$. **Equation:** $HCl + NaOH rightarrowNaCl + H\_2O$ (Mole ratio is 1:1) **Setup:** $frac0.10times21.01 = fracM\_2times25.01$ (p. 130) **Solve for M_2:** $M\_2 = frac0.10times21.025.0 = 0.084\tilde{m}ol/dm^3$ (p. 130)

8) Required practicals / experiments

Preparation of Potassium Nitrate by Titration (p. 128)

- **Aim:** To prepare a pure, dry sample of a soluble salt (potassium nitrate) from an acid and an alkali.
- **Apparatus:** Burette, clamp stand, $25\tilde{c}m^3$ pipette, safety filler, conical flask, evaporating basin, Bunsen burner, tripod, gauze (p. 127).
- **Method:**
  i. Use a pipette and safety filler to measure exactly $25.0\tilde{c}m^3$ of potassium hydroxide solution into a conical flask (p. 128).
  ii. Add 4 drops of thymolphthalein indicator. The solution will turn blue (p. 128).
  iii. Fill a burette with dilute nitric acid, ensuring the jet is full, and record the initial volume (p. 128).
  iv. Slowly add the acid from the burette to the flask, swirling constantly, until the blue colour just disappears. This is the end-point (p. 128). Record the final volume and calculate the volume of acid added.
  v. Repeat the experiment using the same volumes of acid and alkali but **without the indicator** to get a pure solution of the salt.
  vi. Pour the resulting potassium nitrate solution into an evaporating basin and heat to evaporate about half the water (p. 128).
  vii. Allow the concentrated solution to cool and crystallise. Filter and dry the crystals (p. 128).
- **Variables:**
  - **Independent Variable (IV):** Volume of nitric acid added.

- **Dependent Variable (DV):** Colour change of the indicator.
  - **Control Variables:** Concentration of acid and alkali, volume of alkali.
- **Safety:** Wear eye protection. Indicator can stain (p. 127).
- **Sources of Error:** Misjudging the end-point colour change; parallax error when reading the burette; not filling the burette jet.
- **Expected Results:** A specific volume of acid will be required for neutralisation. On cooling the final solution, white crystals of potassium nitrate will form.

9) Data handling and graphing

This chapter primarily uses data from **titration experiments**.

- **Tables:** Titration results are recorded in a table with columns for Initial Burette Reading (cm^3), Final Burette Reading (cm^3), and Titre/Volume Added (cm^3). A row is used for a rough titration and then several accurate ones.
- **Choosing Data:** Consistent results (titres that are within $0.10\tilde{c}m^3$ of each other) should be used to calculate a mean average (p. 129). Anomalous results (those that are not close to the others) should be ignored.
- **Calculations:** The mean titre is used in the titration formula ($M\_1V\_1/n\_1 = M\_2V\_2/n\_2$) to calculate an unknown concentration (p. 130).
- **Typical Exam Prompts:** "Complete the table by calculating the titre for each experiment." "Identify the two best results and calculate the average volume of acid used." "Use your average volume to calculate the concentration of the alkali."

10) Common misconceptions and exam tips

- **Misconception:** Strong acid means it's a concentrated acid.
  - **Correct Understanding: Strength** refers to how much an acid dissociates into ions, while **concentration** refers to how much acid is dissolved in water (p. 119). You can have a dilute strong acid.
  - **Quick Tip:** Think of 'strength' as the *quality* of the acid and 'concentration' as the *quantity*.
- **Misconception:** All bases are alkalis.
  - **Correct Understanding:** An alkali is a base that is **soluble** in water (p. 122). Many bases, like copper(II) oxide, are insoluble.
  - **Quick Tip:** All alkalis are bases, but not all bases are alkalis.
- **Misconception:** In precipitation, the solid is just filtered off and is ready.
  - **Correct Understanding:** The precipitate must be **washed with distilled water** after filtering to remove any soluble impurities before it is dried (p. 124).
  - **Quick Tip:** Remember the steps: Mix → Filter → Wash → Dry.
- **Misconception:** Forgetting state symbols in equations.

- **Correct Understanding:** State symbols (s), (l), (g), (aq) are crucial, especially for showing a precipitate (s) or dissolved ions (aq).
- **Quick Tip:** For precipitation, the ionic equation is always: `cation(aq) + anion(aq) → solid_salt(s)`.

11) Exam-style practice

Multiple Choice Questions (MCQs)

1. Which of the following pH values represents the strongest acid?
   A) 6
   B) 4
   C) 1
   D) 7
   (Answer: C. The lower the pH below 7, the stronger the acid.)
2. A weak acid is best described as one that...
   A) is very dilute.
   B) does not react with metals.
   C) is only partially ionised in water.
   D) has a pH greater than 7.
   (Answer: C. Weakness refers to the degree of dissociation/ionisation.)
3. Which of the following salts is insoluble in water?
   A) Sodium nitrate
   B) Potassium chloride
   C) Silver chloride
   D) Ammonium sulfate
   (Answer: C. All nitrates, potassium salts, and ammonium salts are soluble; silver chloride is a key exception to the chloride rule.)
4. The reaction $H^+(aq) + OH^-(aq)$
   $rightarrow H\_2O(l)$ is the ionic equation for...
   A) Precipitation
   B) Neutralisation
   C) Dissociation
   D) Reduction
   (Answer: B. This shows the hydrogen ion from an acid reacting with the hydroxide ion from an alkali to form water.)
5. What gas is produced when an acid reacts with a metal carbonate?
   A) Hydrogen
   B) Oxygen
   C) Carbon dioxide
   D) Chlorine
   (Answer: C. Acid + Carbonate → Salt + Water + Carbon Dioxide.)

6. Which method is most suitable for preparing copper(II) sulfate from copper(II) oxide?
   A) Acid + excess insoluble base
   B) Titration
   C) Precipitation
   D) Acid + metal
   (Answer: A. Copper(II) oxide is an insoluble base.)
7. To prepare the insoluble salt lead(II) iodide, you should mix aqueous solutions of...
   A) Lead(II) sulfate and sodium iodide
   B) Lead(II) nitrate and potassium iodide
   C) Lead metal and iodine
   D) Hydrochloric acid and lead(II) nitrate
   (Answer: B. Both lead(II) nitrate and potassium iodide are soluble salts that provide the required ions.)
8. The term for a salt that contains water molecules within its crystal structure is...
   A) Anhydrous
   B) Neutral
   C) Hydrated
   D) Saturated
   (Answer: C. A hydrated salt contains water of crystallisation.)
9. What is the correct test for a sulfate ion?
   A) Add nitric acid and silver nitrate solution.
   B) Add hydrochloric acid and barium chloride solution.
   C) Add sodium hydroxide and warm.
   D) Add any acid and test gas with limewater.
   (Answer: B. This forms a white precipitate of barium sulfate.)
10. In a titration, a pipette is used to...
   A) add the acid in small, measured drops.
   B) measure an accurate and fixed volume of the alkali.
   C) see the colour change clearly.
   D) hold the conical flask.
   (Answer: B. A pipette is designed to deliver a precise, fixed volume.)

Short-Answer Questions

1. **Distinguish** between a strong acid and a concentrated acid.
   - **Marking points:** A strong acid is one that fully dissociates into its ions in water (1). Concentration refers to the number of moles of acid dissolved per unit volume of water (1). An acid can be strong but dilute, or weak but concentrated (1).
2. Write the word, chemical, and ionic equations for the reaction between dilute nitric acid and potassium hydroxide.
   - **Marking points:**
     - **Word:** Nitric acid + Potassium hydroxide → Potassium nitrate + Water (1).
     - **Chemical:** $HNO\_3(aq) + KOH(aq)$ $rightarrow KNO\_3(aq) + H\_2O(l)$ (1).

- **Ionic:** $H^+(aq) + OH^-(aq) rightarrow H\_2O(l)$ (1).

3. **Describe** the steps to prepare a pure, dry sample of the insoluble salt lead(II) sulfate.
   - **Marking points:** Mix a solution of a soluble lead salt (e.g., lead(II) nitrate) with a solution of a soluble sulfate (e.g., sodium sulfate) (1). Filter the mixture to collect the precipitate (1). Wash the precipitate with distilled water (1). Dry the precipitate in a warm oven or between filter paper (1).
4. A student tests an unknown solution. Adding dilute nitric acid followed by silver nitrate solution produces a cream precipitate. **Identify** the ion present and write an ionic equation for the reaction.
   - **Marking points:** The ion is bromide / Br^- (1). The ionic equation is $Ag^+(aq) + Br^-(aq) rightarrow AgBr(s)$ (1).
5. Blue crystals of hydrated copper(II) sulfate are heated. **State** the colour change observed and **name** the white solid product.
   - **Marking points:** The colour changes from blue to white (1). The product is anhydrous copper(II) sulfate (1).

Structured Questions

Question 1

A student wants to find the concentration of a solution of sodium hydroxide. She titrates $25.0\tilde{c}m^3$ of the sodium hydroxide solution with $0.100\tilde{m}ol/dm^3$ hydrochloric acid. Her results are shown in the table.

Titration	Initial burette reading / cm^3	Final burette reading / cm^3	Titre / cm^3
1 (Rough)	0.50	22.80	22.30
2	1.20	23.10	
3	23.10	44.90	21.80

a) **Complete** the table by calculating the titre for titration 2. (1 mark)
b) **Identify** the two best results and **calculate** the average titre. (2 marks)
c) The balanced equation is: $HCl(aq) + NaOH(aq) rightarrow NaCl(aq) + H\_2O(l)$. **Use** your average titre from (b) to **calculate** the concentration of the sodium hydroxide solution in mol/dm^3. (3 marks)

- **Marking Points:**
  - a) Titre 2 = 23.10 - 1.20 = **21.90** cm^3 (1).
  - b) Best results are Titration 2 (21.90) and Titration 3 (21.80) as they are concordant (1). Average titre = (21.90 + 21.80) / 2 = **21.85** cm^3 (1).
  - c) Moles of HCl = $(21.85/1000) times 0.100 = 0.002185\tilde{m}ol$ (1). Mole ratio HCl:NaOH is 1:1, so moles of NaOH = $0.002185\tilde{m}ol$ (1). Concentration of NaOH = moles / volume in dm^3 = $0.002185/(25.0/1000) = textbf{0.0874} mol/dm^3$ (1).

Question 2
Salts can be prepared by different methods.
a) **Name** the method used to prepare potassium chloride from potassium hydroxide. **Explain** why an indicator is needed for the first part of the experiment. (2 marks)
b) **Describe** how to prepare zinc chloride crystals starting with zinc carbonate and an acid. (4 marks)
c) **Write** an ionic equation, including state symbols, for the preparation of zinc chloride in (b). (2 marks)

- **Marking Points:**
    - a) Method is **titration** (1). An indicator is needed because both reactants (acid and alkali) are colourless solutions, so it shows when the neutralisation point (end-point) is reached (1).
    - b) Add excess zinc carbonate to hydrochloric acid (1). Stir and wait for the effervescence to stop (1). Filter to remove the excess unreacted zinc carbonate (1). Evaporate the filtrate until saturated, then cool to form crystals (1).
    - c) $ZnCO\_3(s) + 2H^{+}(aq)$ $rightarrow Zn^{2+}(aq) + H\_2O(l) + CO\_2(g)$ (1 for species, 1 for state symbols).

12) Quick revision checklist

- [ ] Can you define an acid as a proton donor and a base as a proton acceptor? (p. 118)
- [ ] Do you know the difference between a strong and weak acid based on dissociation? (p. 118-119)
- [ ] Can you explain that strength and concentration are not the same thing? (p. 119)
- [ ] Can you state the colours of litmus and universal indicator in acidic, neutral, and alkaline conditions? (p. 117)
- [ ] Do you know that the ionic equation for neutralisation is $H^{+} + OH^{-}$ $rightarrow H\_2O$? (p. 120)
- [ ] Can you name the salts formed from hydrochloric, sulfuric, and nitric acids? (p. 121)
- [ ] Can you recall the general solubility rules for salts? (p. 120)
- [ ] Can you describe the method to make a soluble salt from an acid and an excess of an insoluble reactant (metal, base, or carbonate)? (p. 121-123)
- [ ] Can you describe how to use titration to make a soluble salt from an acid and an alkali? (p. 122)
- [ ] Can you describe how to make an insoluble salt by precipitation? (p. 124)
- [ ] Can you describe the chemical tests for carbonate, sulfate, chloride, and nitrate ions? (p. 124-125)
- [ ] Do you know what 'water of crystallisation' means? (p. 126)
- [ ] Can you describe the test for water using anhydrous copper(II) sulfate? (p. 127)

13) Flashcards (ready-to-use)

Q1: What is the definition of a strong acid?
A1: An acid that is completely dissociated or ionised in aqueous solution (p. 119).

Q2: What is the pH range for acids, alkalis, and neutral substances?

A2: Acids: pH < 7. Neutral: pH = 7. Alkalis: pH > 7 (p. 117).

Q3: What is an alkali?

A3: A base that is soluble in water (p. 122).

Q4: What are the products of a neutralisation reaction?

A4: A salt and water (p. 120).

Q5: State the ionic equation for the reaction between any strong acid and strong alkali.

A5: $H^+(aq) + OH^-(aq)$
$rightarrowH\_2O(l)$ (p. 120).

Q6: Name the salt produced from sulfuric acid and zinc oxide.

A6: Zinc sulfate (p. 121).

Q7: Which two types of salt are always soluble?

A7: All sodium, potassium and ammonium salts, and all nitrates (p. 120).

Q8: What is the name of the method used to make an insoluble salt?

A8: Precipitation (p. 124).

Q9: Why is excess reactant used when making a soluble salt from an insoluble base?

A9: To ensure all the acid is completely reacted and neutralised (p. 121).

Q10: How is the excess reactant removed in this method?

A10: By filtration (p. 121).

Q11: Which piece of apparatus is used to accurately measure a fixed volume of alkali for a titration?

A11: A pipette (with a safety filler) (p. 127).

Q12: What is the chemical test for a carbonate ion?

A12: Add dilute acid; it will effervesce (fizz). The gas produced (carbon dioxide) turns limewater milky (p. 125).

Q13: What is observed when silver nitrate is added to a solution containing chloride ions?

A13: A white precipitate of silver chloride forms (after adding nitric acid) (p. 125).

Q14: What is 'water of crystallisation'?

A14: Water molecules that are chemically bonded into the structure of a salt crystal (p. 126).

Q15: How can you test for the presence of water?

A15: Add it to white anhydrous copper(II) sulfate. It will turn blue (p. 127).

14) 60-second recap

This chapter covers acids, bases, and salts. Acids are proton donors with a pH below 7, while bases are proton acceptors. Strong acids fully dissociate in water; weak acids only partially. We use indicators to measure pH. Neutralisation is the reaction between an acid and a base, forming a salt and water. There are four ways to make soluble salts: acid plus a metal, insoluble base, or carbonate, or by titration for soluble bases. Insoluble salts are made by precipitation. You must know the solubility rules and the chemical tests for common anions like carbonates, sulfates, and halides. Finally, some salts are hydrated, meaning they contain water of crystallisation, which can be driven off by heating.

15) References to pages

16) Excluded "Going further" sections (not summarized)

Section title	Pages
Calculation of water of crystallisation	p. 127
Total excluded:	**1**