LEARNERS' ALLEY
Designed by Khow Willard

Sec 3

IP Chemistry

Book 2

For the 6092 syllabus



NAME _____

Sec 3 End-of-Year Examination

IP Chemistry, Secondary 3. There are two papers and no practical paper. The Sec 4 O Level has three papers (see the Sec 4 book).



Paper 1
Multiple Choice

DURATION	QUESTIONS	WEIGHTING	TIME PER QUESTION
45 min	30	39.6%	1.5 min

- 30 compulsory multiple choice questions.

Paper 2
Structured Questions

DURATION	MARKS	WEIGHTING	TIME PER MARK
1 h 15 min	55	60.4%	1.4 min

- **Section A** (45 marks): compulsory structured questions.
- **Section B** (10 marks): two questions, answer **one**.

No Paper 3
Practical

- There is no practical paper in the Sec 3 end-of-year examination.
- The Sec 4 O Level has a practical paper (Paper 3). It is in the Sec 4 book.

How to use these notes

Fill in the blanks

Write the missing key word on each line, e.g. electrons are added to the _____. The answers are the phrases examiners look for.

Labels

NOTE marks a special case, an exception or a point that is easy to miss. EXAM TIP marks advice on how to phrase an answer or set out working. ANSWER STRUCTURE shows how to set out the answer to a type of question you will meet again and again.

Explanations and examples

Numbered points in an explanation are separate mark points: write every one. Each WORKED EXAMPLE sets out the steps one at a time.

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08 Acids and Bases



1 Acids

1.1 What an acid is

DEFINITION An **acid** is a substance that produces _____, in aqueous solution.

Acid	Formula	Ions in aqueous solution
hydrochloric acid	HCl	H^+ and Cl^-
nitric acid	HNO_3	H^+ and NO_3^-
sulfuric acid	H_2SO_4	H^+ and SO_4^{2-}
ethanoic acid	CH_3COOH	H^+ and CH_3COO^-

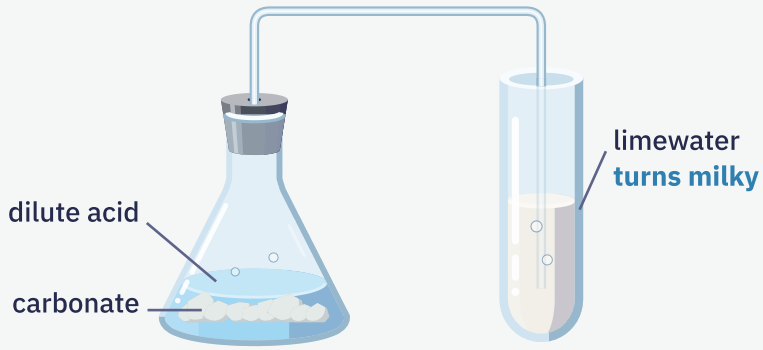
- Acids ionise only when dissolved in water. A solid acid such as citric acid does not conduct electricity, but its _____ solution does, because it contains mobile ions.
- Acids taste sour, turn blue litmus paper _____ and have a pH below 7.

1.2 Reactions of acids

Acid reacts with	Products	Example
a reactive **metal**	salt + _____	$Mg(s) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + H_2(g)$
a **base**	salt + water	$MgO(s) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + H_2O(l)$
a **carbonate**	salt + water + _____	$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$



hydrogen extinguishes a lighted splint with a pop



carbon dioxide turns limewater milky

- Sulfuric acid forms **sulfates**, nitric acid forms **nitrates** and hydrochloric acid forms **chlorides**.
- Unreactive metals such as copper and silver do not react with dilute acids.

NOTE Lead seems not to react with hydrochloric acid or sulfuric acid. The first reaction forms insoluble lead(II) chloride or lead(II) sulfate, which coats the lead and stops further reaction.

1 Acids (continued)

1.3 Strong and weak acids

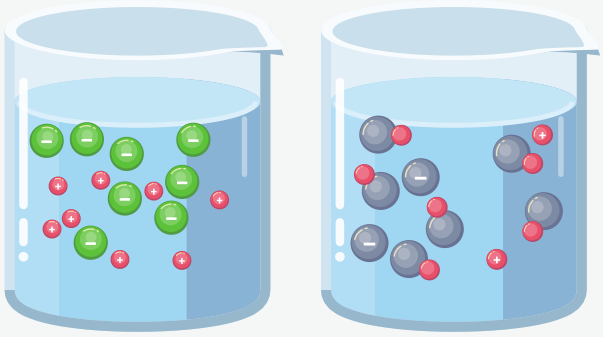
DEFINITION A **strong acid** is _____ ionised in aqueous solution, e.g. hydrochloric, nitric and sulfuric acids.

DEFINITION A **weak acid** is only _____ ionised in aqueous solution, e.g. ethanoic acid and citric acid.

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

$CH_3COOH(aq) \rightleftharpoons CH_3COO^-(aq) + H^+(aq)$

- The ⇌ sign shows a reversible reaction: the ions keep re-forming molecules.
- **Strength** is how much of the acid ionises. **Concentration** is how much acid is dissolved in a given volume of water. A strong acid can be dilute, and a weak acid can be concentrated.



hydrochloric acid: all ions

ethanoic acid: mostly molecules

2 Bases and Alkalis

2.1 Bases and alkalis

DEFINITION A **base** is a metal oxide or metal hydroxide. An **alkali** is a base that is _____ in water and produces _____, in aqueous solution.

Bases that do not dissolve in water

CuO, ZnO, $Mg(OH)_2$ and most other metal oxides and hydroxides

Alkalis: bases that dissolve

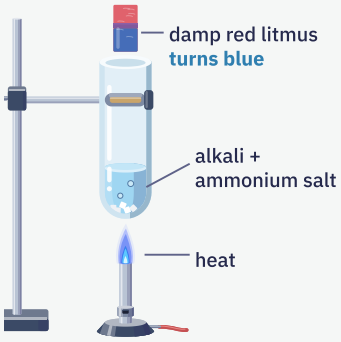
$NaOH$, KOH, $Ca(OH)_2$ and $NH_3(aq)$, clear solutions containing OH^- ions

- Alkalis taste bitter, feel soapy, turn red litmus paper _____ and have a pH above 7.
- Sodium hydroxide is a **strong** alkali: it is completely ionised. Aqueous ammonia is a **weak** alkali: $NH_3(g) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$.

2.2 Reactions of bases

Neutralisation is the reaction between hydrogen ions and hydroxide ions to produce water:
$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

Base reacts with	Products	Example
an **acid**	salt + _____	$2NaOH(aq) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$
an **ammonium salt** (alkalis, on heating)	salt + water + _____	$KOH(aq) + NH_4Cl(s) \rightarrow KCl(aq) + H_2O(l) + NH_3(g)$



- Ammonia has a pungent smell. It turns **damp red litmus paper blue**.
- An insoluble base reacts with the H^+ ions of an acid, e.g. $CuO(s) + 2H^+(aq) \rightarrow Cu^{2+}(aq) + H_2O(l)$.

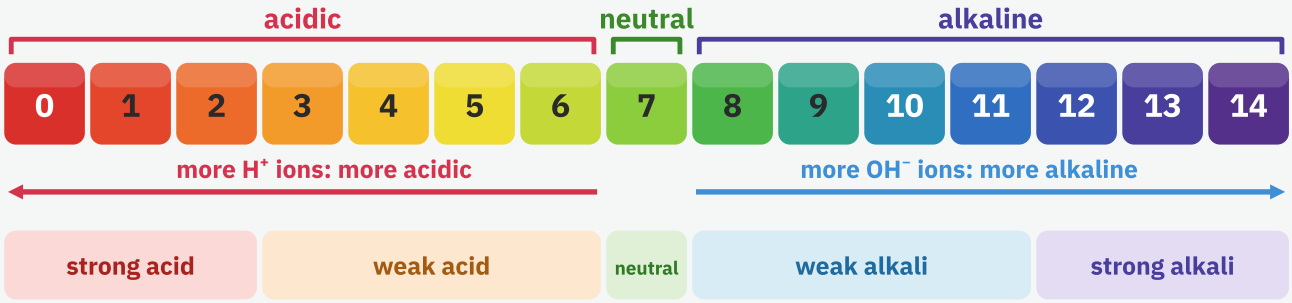
WORKED EXAMPLE Milk of magnesia contains magnesium hydroxide. Write the equation for its reaction with hydrochloric acid in the stomach.

Base + acid → salt + water. The salt is magnesium chloride, $MgCl_2$.

$Mg(OH)_2(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + 2H_2O(l)$

3 The pH Scale

3.1 pH and Universal Indicator



pH	Ion concentrations	Solution	Universal Indicator
below 7	more H^+ than OH^-	_____	red, orange or yellow
7	H^+ and OH^- equal	neutral	_____
above 7	more OH^- than H^+	_____	blue, then violet

- The _____ the pH, the higher the concentration of H^+ ions. A solution of pH 4 has 1000 times more H^+ ions than one of pH 7.
- pH is measured with Universal Indicator (paper or solution), a pH meter, or a pH sensor connected to a data logger.

3.2 Other indicators

Indicator	Colour in acid	Changes at pH	Colour in alkali
methyl orange	red	3 to 5	yellow
screened methyl orange	violet	3 to 5	green
litmus	red	5 to 8	blue
thymolphthalein	colourless	9 to 10.5	blue

3.3 Controlling the pH of soil

- Most plants grow best in soil that is neutral or slightly acidic. Soil that is too acidic or too alkaline reduces the nutrients a plant can absorb.
- Acid rain, fertilisers and microbes in the soil can make it too acidic.
- Farmers add **slaked lime**, calcium hydroxide, to _____ the excess acid. Too much makes the soil too alkaline.

4 Oxides

4.1 Four types of oxide

Type	Formed by	Reacts with	Examples
basic	most _____	acids, forming a salt and water	CuO, CaO, MgO, Na_2O
amphoteric	some metals	_____ acids and alkalis, forming a salt and water	ZnO, Al_2O_3, PbO
acidic	most _____	alkalis, forming a salt and water	CO_2, SO_2, SO_3, P_4O_{10}, SiO_2
neutral	some non-metals	_____ acids nor alkalis	H_2O, CO, NO

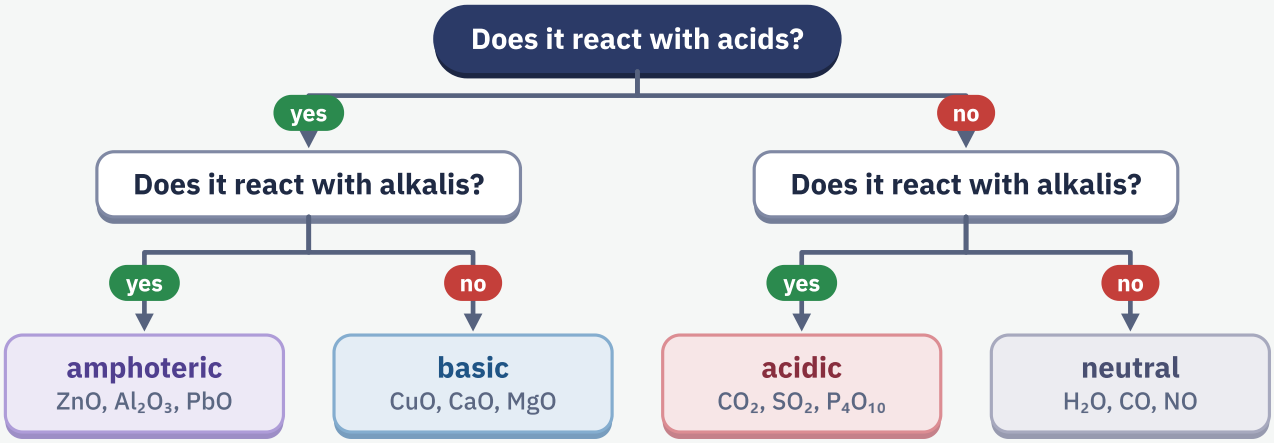
- Most basic oxides are insoluble solids. Sodium oxide and potassium oxide dissolve in water to form alkalis.
- Most acidic oxides dissolve in water to form acids: $SO_2(g) + H_2O(l) \rightarrow H_2SO_3(aq)$.

$ZnO(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2O(l)$

$ZnO(s) + 2NaOH(aq) \rightarrow Na_2ZnO_2(aq) + H_2O(l)$

$CO_2(g) + 2NaOH(aq) \rightarrow Na_2CO_3(aq) + H_2O(l)$

4.2 Classifying an unknown oxide



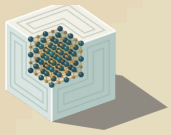
WORKED EXAMPLE Oxide X is a white solid that is insoluble in water. It reacts with dilute hydrochloric acid and with aqueous sodium hydroxide, and forms a colourless solution with each. Classify X and suggest its identity.

X reacts with both an acid and an alkali, so X is **amphoteric**.

It is a white solid oxide of a metal, so X could be **zinc oxide** or **aluminium oxide**.

EXAM TIP Classify an oxide from the element: a metal oxide is basic (or amphoteric for zinc, aluminium and lead) and a non-metal oxide is acidic (or neutral for water, carbon monoxide and nitrogen monoxide).

09 Salts



1 Salts and Solubility

1.1 What a salt is

A **salt** is an ionic compound made of a _____ (positive ion) and an _____ (negative ion).

- A salt forms when the hydrogen ions of an acid are replaced by metal ions or ammonium ions, e.g. in the reactions of acids with metals, bases and carbonates.
- Sulfuric acid gives **sulfates**, nitric acid gives **nitrates** and hydrochloric acid gives **chlorides**.

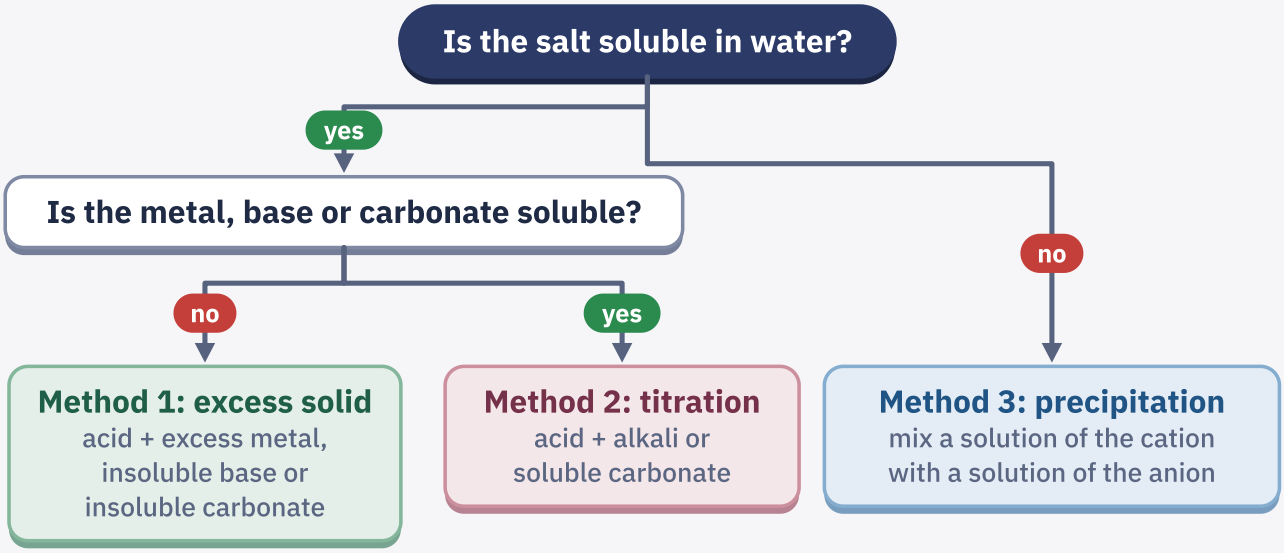
1.2 Solubility rules

Salts	Soluble in water	Insoluble in water
sodium, potassium and ammonium salts	_____	none
nitrates	all	none
chlorides	most	_____ AgCl and _____ $PbCl_2$
sulfates	most	_____ $BaSO_4$, _____ $PbSO_4$ and calcium sulfate $CaSO_4$ (sparingly soluble)
carbonates	sodium, potassium and ammonium carbonates only	_____, e.g. $CaCO_3$, $CuCO_3$, $ZnCO_3$
hydroxides	sodium, potassium and ammonium hydroxides (calcium hydroxide is slightly soluble)	all others, e.g. $Cu(OH)_2$, $Fe(OH)_3$

EXAM TIP Check the solubility of the salt **and** of each starting material before you choose a method to prepare a salt.

2 Preparing Salts

2.1 Choosing a method

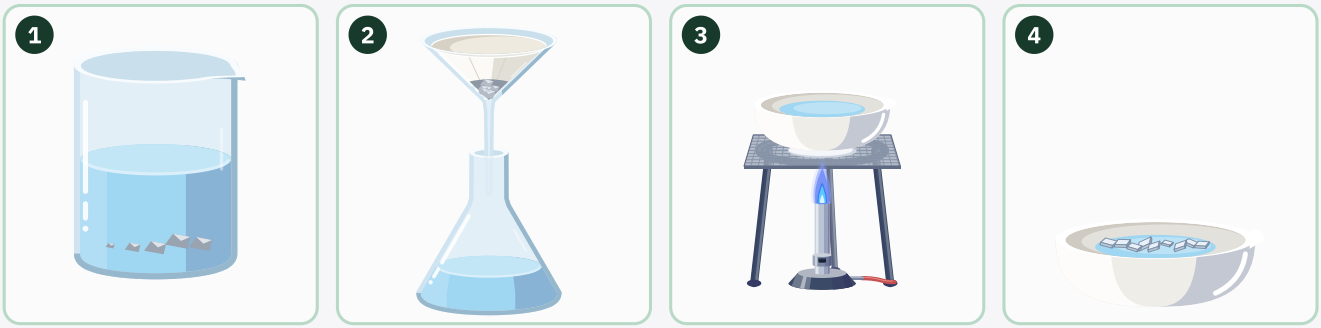


2.2 Method 1: an acid and an excess insoluble substance

acid + metal → salt + hydrogen

acid + insoluble base → salt + water

acid + insoluble carbonate → salt + water + carbon dioxide



add excess zinc to the acid | filter off the excess zinc | heat the filtrate until saturated | cool to crystallise, then filter, wash and dry

- Add the solid in _____, so that all the acid is used up and the salt is not contaminated with acid. The excess solid is insoluble, so it is removed by **filtration**.
- Heat the filtrate until _____: a glass rod dipped in and removed shows small crystals forming as it cools.
- Do not use very reactive metals (they react violently) or unreactive metals such as copper (they do not react with dilute acids). Use the metal oxide or carbonate instead.

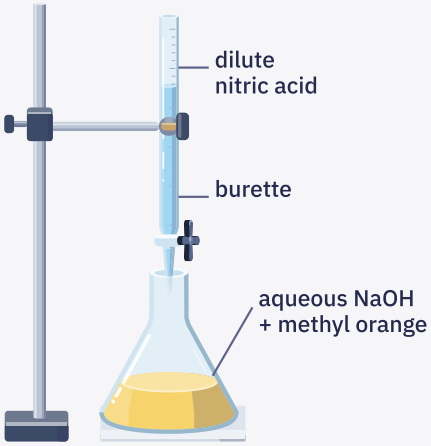
$H_2SO_4(aq) + Zn(s) \rightarrow ZnSO_4(aq) + H_2(g)$

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

$2HCl(aq) + MgCO_3(s) \rightarrow MgCl_2(aq) + CO_2(g) + H_2O(l)$

2 Preparing Salts (continued)

2.3 Method 2: titration



acid + alkali → salt + water

acid + soluble carbonate → salt + water + carbon dioxide

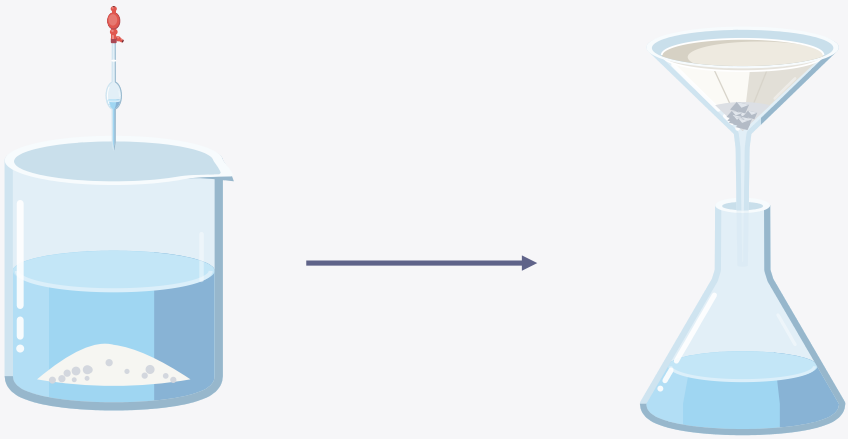
- Used for **sodium, potassium and ammonium** salts, where all the starting materials are soluble.
- An excess of either reactant cannot be filtered off, so the volumes must be _____.

1. Pipette 25.0 cm^3 of the alkali into a conical flask and add a few drops of indicator, e.g. methyl orange.
2. Add the acid from a burette until the indicator just changes colour (the **end-point**). Record the volume of acid used.
3. Repeat with the same volumes but _____, which would make the salt impure.
4. Heat the solution until saturated, cool to crystallise, then filter, wash with a little cold distilled water and dry the crystals.

$HNO_3(aq) + NaOH(aq) \rightarrow NaNO_3(aq) + H_2O(l)$

2.4 Method 3: precipitation

An **insoluble** salt is made by mixing two solutions: one contains its _____ and the other contains its _____. The salt forms as a precipitate.



mix: a white precipitate forms **filter: the precipitate is the residue**

1. Mix the two solutions and stir until no more precipitate forms.
2. Filter to collect the precipitate as the residue.
3. Wash it with _____ to remove soluble impurities, then dry it between sheets of filter paper.

$Ba(NO_3)_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2NaNO_3(aq)$

ionic equation: $Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$

2 Preparing Salts (continued)

2.5 Planning a preparation

Salt	Soluble?	Method	Starting materials
zinc sulfate	yes	excess insoluble substance	dilute sulfuric acid + excess zinc, zinc oxide or zinc carbonate
copper(II) nitrate	yes	excess insoluble substance	dilute nitric acid + excess copper(II) oxide or carbonate (not copper)
sodium nitrate	yes	titration	dilute nitric acid + aqueous sodium hydroxide or sodium carbonate (not sodium)
barium sulfate	_____	precipitation	aqueous barium nitrate + dilute sulfuric acid or any soluble sulfate

WORKED EXAMPLE Describe how you would prepare lead(II) chloride from lead(II) carbonate, dilute nitric acid and aqueous potassium chloride.

Lead(II) chloride is insoluble, so it is made by **precipitation**. Lead(II) carbonate is also insoluble, so first turn it into a soluble lead(II) salt.

Add **excess** lead(II) carbonate to dilute nitric acid: $PbCO_3(s) + 2HNO_3(aq) \rightarrow Pb(NO_3)_2(aq) + CO_2(g) + H_2O(l)$.

Filter off the unreacted lead(II) carbonate. The filtrate is aqueous lead(II) nitrate.

Add excess aqueous potassium chloride to the filtrate: $Pb(NO_3)_2(aq) + 2KCl(aq) \rightarrow PbCl_2(s) + 2KNO_3(aq)$.

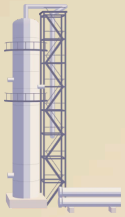
Filter to collect the lead(II) chloride, wash it with distilled water and dry it between sheets of filter paper.

EXAM TIP In a description, name the reagents, say which one is in excess and why, and give every step: filter, heat to saturation, cool, filter, wash and dry.

2.6 Common mistakes

Mistake	Why it is wrong
Using sodium or potassium metal with an acid	Very reactive metals react _____ with acids.
Using copper or silver with a dilute acid	Unreactive metals do _____ with dilute acids.
Not adding the solid in excess (method 1)	Some acid is left, so the salt is _____ with acid.
Leaving the indicator in (method 2)	The indicator makes the salt _____.
Using method 1 for a sodium salt	Sodium compounds are soluble, so an excess cannot be _____ off.
Washing crystals with a lot of water	Use a little **cold** distilled water, or the crystals dissolve.

10 Ammonia



1 The Haber Process

1.1 Making ammonia

- Ammonia is made industrially by the **Haber process**. About 80% of it is used to make _____. It is also used to make plastics, explosives, dyes and cleaning products.

Raw material	Obtained from
nitrogen, N_2	the _____
hydrogen, H_2	the _____ of crude oil fractions

$$N_2(g) + 3H_2(g) _____ 2NH_3(g)$$

Forward reaction

$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$

Backward reaction

$2NH_3(g) \rightarrow N_2(g) + 3H_2(g)$

- The reaction is **reversible**, so some of the ammonia made breaks down again into nitrogen and hydrogen.

NOTE At room temperature nitrogen is very unreactive. Recall from Chemical Bonding that a nitrogen molecule has a strong **triple bond**, N≡N. So the forward reaction needs a high pressure, a relatively high temperature and a catalyst.

1.2 Choosing the conditions

Condition	Effect	Used
pressure	A higher pressure gives a _____ and a faster reaction. But keeping a high pressure is costly: it needs expensive equipment and a large amount of electricity.	200 to 250 atm (a compromise)
temperature	A lower temperature gives a _____, because less ammonia decomposes. But the reaction would be too slow.	450 °C (a compromise)
catalyst	The reaction is still slow under these conditions, so a catalyst speeds it up. It does not change the yield.	finely divided _____

EXAM TIP When asked why a condition is a compromise, give both sides: what a change would do to the **yield**, and what it would do to the **rate** or the **cost**.

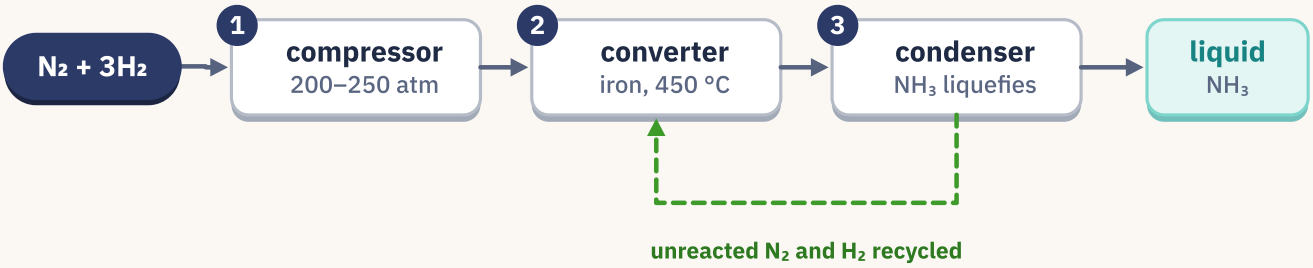
WORKED EXAMPLE Some plants run the Haber process at 300 atm instead of 200 atm. Suggest one advantage and one disadvantage of the higher pressure.

Advantage: a higher yield of ammonia *or* a faster reaction.

Disadvantage: keeping the higher pressure costs more, because more energy is needed.

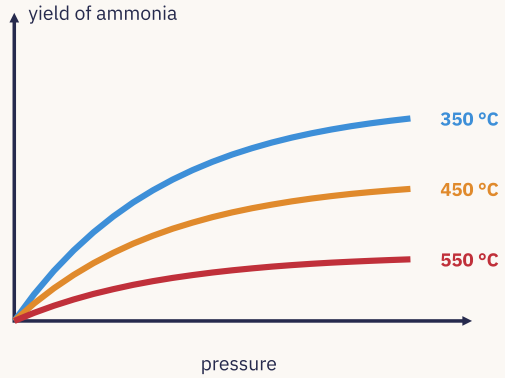
1 The Haber Process (continued)

1.3 Inside the plant



1. Nitrogen and hydrogen are mixed in the ratio **1 : 3** by volume.
2. The mixture is compressed to 200 to 250 atm.
3. The gases flow over the iron catalyst at 450 °C. Only about _____ of the mixture leaving the converter is ammonia.
4. The mixture is cooled. Ammonia _____ into a liquid and is pumped into tanks.
5. The unreacted nitrogen and hydrogen are _____ back to the converter, so none is wasted.

1.4 Reading yield data



- The yield _____ as the pressure increases.
- The yield _____ as the temperature increases.
- The conditions used are a compromise between a good yield, a fast enough rate and a reasonable cost.

WORKED EXAMPLE In the Haber process, a temperature of 450 °C is used instead of 150 °C or 650 °C. Explain why.

1. At 150 °C the yield would be higher, but the reaction would be too **slow**.
2. At 650 °C the reaction is faster, but more ammonia decomposes, so the **yield is lower**.
3. 450 °C is a **compromise**: a reasonable yield at a fast enough rate.

2 Reversible Reactions

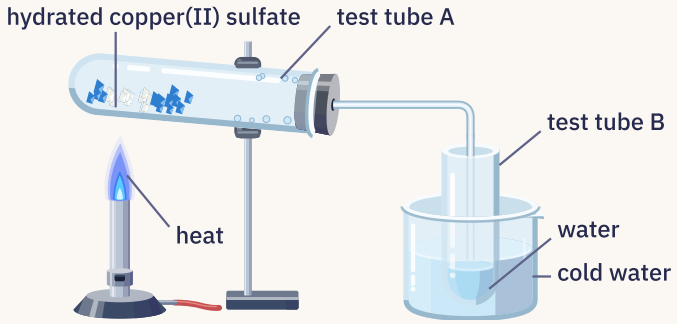
2.1 Reversible and irreversible reactions

DEFINITION A **reversible reaction** can go both _____. Its equation uses the sign ⇌.

	Reversible reaction	Irreversible reaction
What happens	The reactants change into the products, and the products can change _____ into the reactants.	The reactants change into the products, but the products cannot change back.
Examples	the Haber process, heating hydrated copper(II) sulfate, heating ammonium chloride	burning charcoal: the ash cannot be turned back into charcoal

NOTE A reversible reaction does not go to **completion**. Even under the best conditions, the forward and backward reactions keep taking place at the same rate, so the yield is always less than 100%.

2.2 Heating hydrated copper(II) sulfate

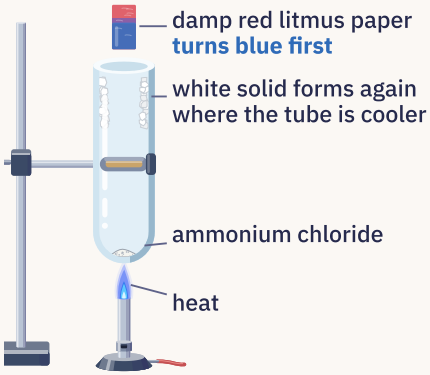


- When heated, blue **hydrated** copper(II) sulfate gives off water and turns into white **anhydrous** copper(II) sulfate. The water is collected in test tube B.
- When the water is added back, the white powder turns _____ again.



$$CuSO_4 \cdot 5H_2O(s) \rightleftharpoons CuSO_4(s) + 5H_2O(l)$$

2.3 Heating ammonium chloride



- When heated, solid ammonium chloride breaks down into **ammonia** and **hydrogen chloride**.
- Where the tube is cooler, the two gases combine to form ammonium chloride _____.
- Ammonia is lighter, so it diffuses faster and turns damp red litmus paper blue before the hydrogen chloride arrives.

$$NH_4Cl(s) \rightleftharpoons NH_3(g) + HCl(g)$$

WORKED EXAMPLE When solid ammonium chloride is heated, it forms ammonia and a gas X. On cooling, X combines with ammonia to form ammonium chloride again. (a) Identify X. (b) Write the equation with state symbols. (c) Ammonia reaches the end of the test tube first. What can you deduce? (d) What is the effect of X on damp blue litmus paper?

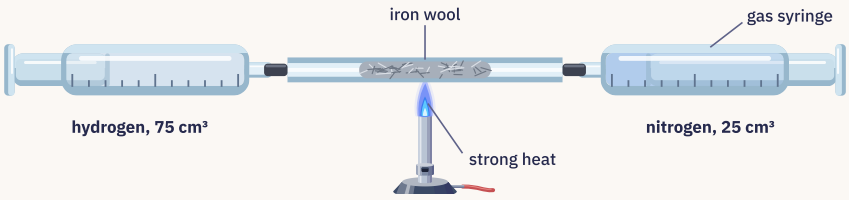
(a) X is **hydrogen chloride**, HCl.

(b) **$NH_4Cl(s) \rightleftharpoons NH_3(g) + HCl(g)$**

(c) Ammonia has a lower relative molecular mass (17) than hydrogen chloride (36.5), so it diffuses **faster**.

(d) Hydrogen chloride turns damp blue litmus paper **red**.

2.4 Yield from a reversible reaction



WORKED EXAMPLE 25 cm^3 of nitrogen and 75 cm^3 of hydrogen are passed forward and backward over heated iron wool until no further reaction is observed. Suggest the volume of ammonia produced. Explain your answer.

1. $N_2 + 3H_2 \rightleftharpoons 2NH_3$, so from the mole ratio, 25 cm^3 of nitrogen and 75 cm^3 of hydrogen could give at most **50 cm^3** of ammonia.
2. The reaction is **reversible**, so it does not go to completion.
3. The volume of ammonia is **less than 50 cm^3**.

NOTE At the same temperature and pressure, volumes of gases react in the same ratio as their moles. The iron wool is a **catalyst**.

2.5 Other industrial processes

WORKED EXAMPLE The first stage of making nitric acid is a reversible reaction in which ammonia reacts with oxygen to form nitrogen monoxide, NO, and water. (a) Write the balanced equation. (b) The yield of NO is higher at lower temperatures, yet 600 °C is used. What would happen below 600 °C?

(a) **$4NH_3 + 5O_2 \rightleftharpoons 4NO + 6H_2O$**

(b) The yield of NO would be higher, but the reaction would be too **slow**.

EXAM TIP For data on another industrial process, such as the Contact process, first describe the trend: how the yield changes as the pressure or the temperature increases, quoting values. Then weigh the yield against the rate and the cost, as for the Haber process.

11 Qualitative Analysis

1 Testing for Cations

1.1 How the tests work

Qualitative analysis identifies the _____ in a substance, and the _____ it gives off, from what is observed when reagents are added.

- Cations are identified with **aqueous sodium hydroxide** and **aqueous ammonia**. Add a few drops, then add the reagent _____.
- Note the _____ of any precipitate, whether it _____, and whether _____ is given off on warming.
- The precipitate is the **metal hydroxide**, which is insoluble in water. Sodium, potassium and ammonium ions give no precipitate.

$Cu^{2+}(aq) + 2OH^-(aq) \rightarrow Cu(OH)_2(s)$

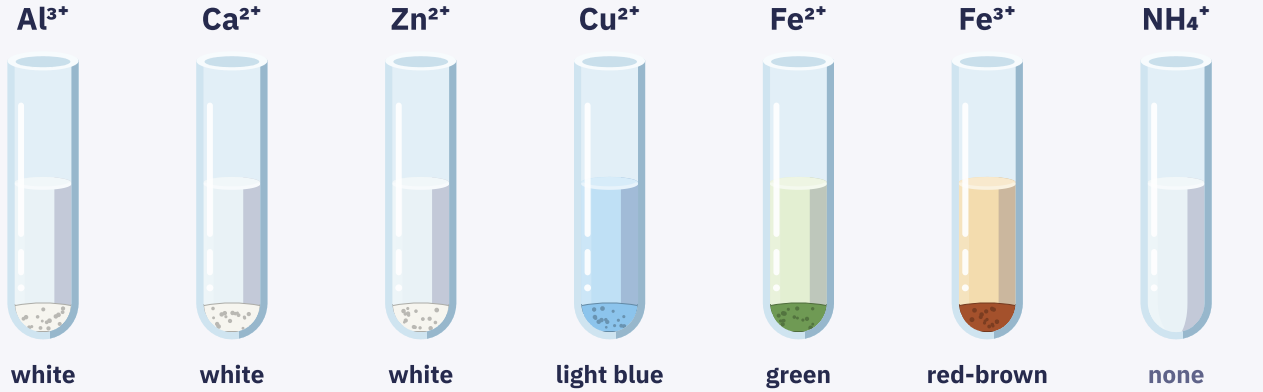
$Fe^{3+}(aq) + 3OH^-(aq) \rightarrow Fe(OH)_3(s)$

1.2 Results for each cation

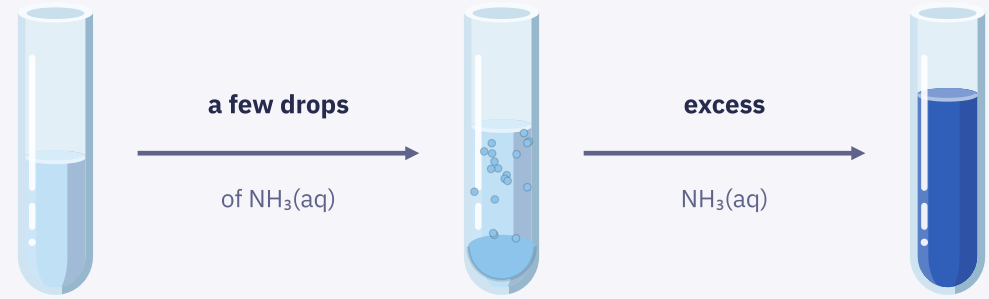
Cation	Aqueous sodium hydroxide	Aqueous ammonia
aluminium, Al^{3+}	white precipitate, _____ in excess giving a colourless solution	white precipitate, _____ in excess
ammonium, NH_4^+	ammonia produced on warming	no change
calcium, Ca^{2+}	white precipitate, insoluble in excess	_____
copper(II), Cu^{2+}	_____ precipitate, insoluble in excess	light blue precipitate, soluble in excess giving a _____
iron(II), Fe^{2+}	_____ precipitate, insoluble in excess	green precipitate, insoluble in excess
iron(III), Fe^{3+}	_____ precipitate, insoluble in excess	red-brown precipitate, insoluble in excess
zinc, Zn^{2+}	white precipitate, soluble in excess giving a colourless solution	white precipitate, _____ in excess giving a colourless solution

1 Testing for Cations (continued)

1.3 The colours of the precipitates



- The green precipitate of iron(II) hydroxide turns _____ on standing.
- In excess aqueous sodium hydroxide, the white precipitates of aluminium and zinc hydroxide _____ to give colourless solutions. Calcium hydroxide does not.
- Only copper(II) hydroxide dissolves in excess aqueous ammonia to give a **dark blue** solution.



1 Testing for Cations (continued)

1.4 Identifying a cation: the complete flowcharts

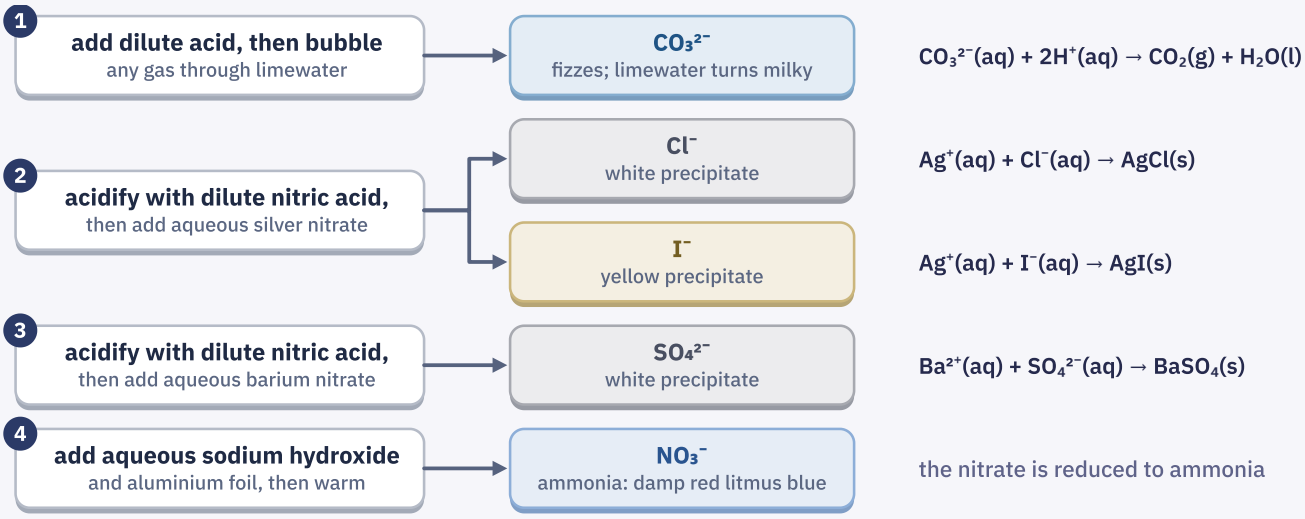


EXAM TIP Describe a result in full: the **colour** of the precipitate and what happens **in excess**, e.g. "white precipitate, soluble in excess giving a colourless solution". Aluminium and zinc give the same result with aqueous sodium hydroxide, so aqueous ammonia is needed to tell them apart.

2 Testing for Anions

2.1 Tests for anions

Anion	Test	Result
carbonate, CO_3^{2-}	add dilute acid, and bubble the gas into limewater	_____: the gas is carbon dioxide, which gives a white precipitate with limewater
chloride, Cl^-	acidify with dilute nitric acid, then add aqueous **silver nitrate**	_____ precipitate of silver chloride
iodide, I^-	acidify with dilute nitric acid, then add aqueous **silver nitrate**	_____ precipitate of silver iodide
sulfate, SO_4^{2-}	acidify with dilute nitric acid, then add aqueous **barium nitrate**	_____ precipitate of barium sulfate
nitrate, NO_3^-	add aqueous sodium hydroxide, then **aluminium** foil, and warm carefully	_____ is produced: it turns damp red litmus paper blue



2.2 Why dilute nitric acid is added first

- The acid removes any _____ (and hydroxide) ions, which would also form a precipitate with silver or barium ions and give a false result.
- Nitric acid is used because all nitrates are soluble, so it adds no ions that form a precipitate.

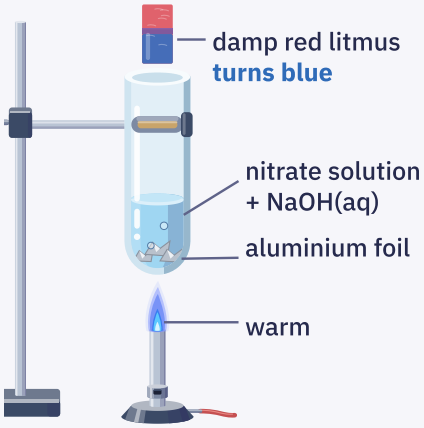
EXAM TIP Do not acidify with hydrochloric acid in the chloride test, or with sulfuric acid in the sulfate test. They add the very ions being tested for.

WORKED EXAMPLE An unknown solution contains either sodium chloride or sodium iodide. Describe how you would find out which anion is present.

1. Acidify a sample with dilute **nitric acid**, then add aqueous **silver nitrate**.
2. A **white** precipitate shows chloride ions. A **yellow** precipitate shows iodide ions.

2 Testing for Anions (continued)

2.3 The nitrate test



- Aluminium _____ the nitrate ions to ammonia, which turns damp red litmus paper blue.
- Ammonium ions also give ammonia when warmed with aqueous sodium hydroxide, even without aluminium. Test for them first.
- Hold the litmus paper at the mouth of the tube. It must not touch the alkali in the tube.

3 Testing for Gases

3.1 Tests for gases



oxygen
glowing splint relights

hydrogen
lighted splint: pop

carbon dioxide
limewater turns milky

ammonia
red litmus turns blue

chlorine
blue litmus bleached

sulfur dioxide
purple to colourless

Gas	Test	Result
oxygen, O_2	insert a **glowing** splint	the splint _____
hydrogen, H_2	hold a **lighted** splint at the mouth	the splint is extinguished with a _____
carbon dioxide, CO_2	bubble the gas into limewater	_____, which dissolves with excess carbon dioxide
ammonia, NH_3	hold damp **red** litmus paper at the mouth	it turns _____
chlorine, Cl_2	hold damp **blue** litmus paper at the mouth	it turns red, then is _____ white
sulfur dioxide, SO_2	bubble into, or hold filter paper soaked in, acidified potassium manganate(VII)	it turns from _____

- Oxygen, hydrogen and carbon dioxide are colourless and odourless. Ammonia, chlorine and sulfur dioxide have a **pungent** smell, and chlorine is **greenish-yellow**.

3 Testing for Gases (continued)

3.2 Putting the tests together

WORKED EXAMPLE A solution of compound K gives these results.
(1) Warmed with aqueous sodium hydroxide: a light blue precipitate, and a gas that turns damp red litmus paper blue.
(2) With aqueous ammonia: a light blue precipitate, soluble in excess to give a dark blue solution.
(3) Acidified with dilute nitric acid, then aqueous barium nitrate added: a white precipitate.
Identify the ions in K.

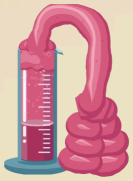
The gas is ammonia, so K contains **ammonium** ions, NH_4^+.

The light blue precipitate that dissolves in excess ammonia to a dark blue solution shows **copper(II)** ions, Cu^{2+}.

The white precipitate with barium nitrate shows **sulfate** ions, SO_4^{2-}.

EXAM TIP When a gas is given off, name it and give its test and result. The tables in this topic are printed in the practical paper (Paper 3) only, so learn them for Papers 1 and 2.

17 Rate of Reactions



1 Measuring the Rate

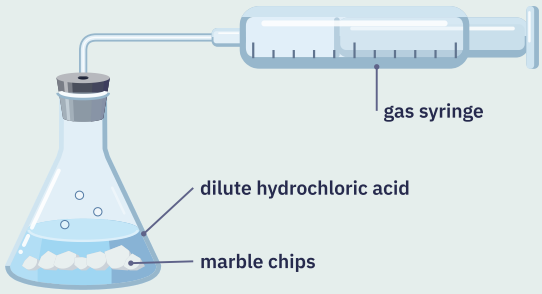
1.1 What the rate is

DEFINITION The **rate of reaction** is how fast a _____ or a _____, measured as the change in amount _____.

- Follow a property that changes as the reaction goes on: the **volume of gas** given off, or the **mass** of the reaction mixture.

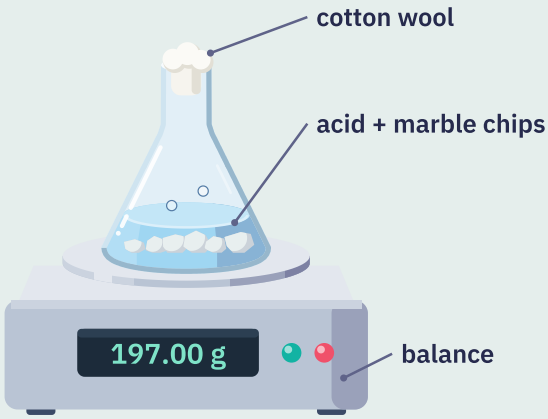
$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$

Volume of gas



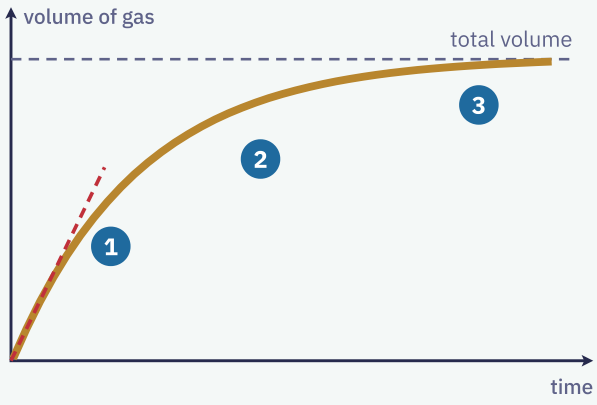
Read the volume in the **gas syringe** at regular time intervals.

Loss in mass



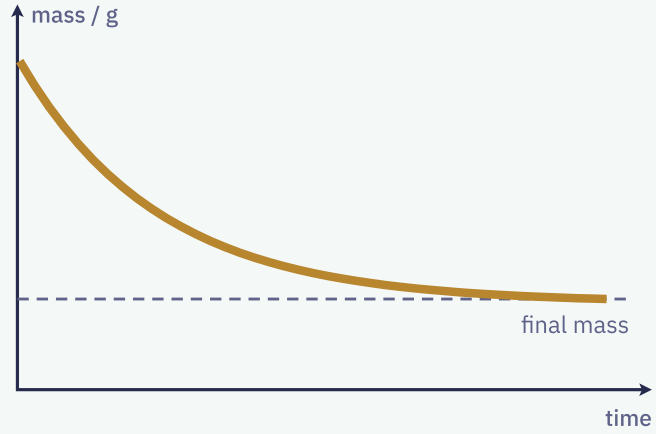
Carbon dioxide escapes, so the mass on the **balance** falls. The cotton wool stops acid spraying out.

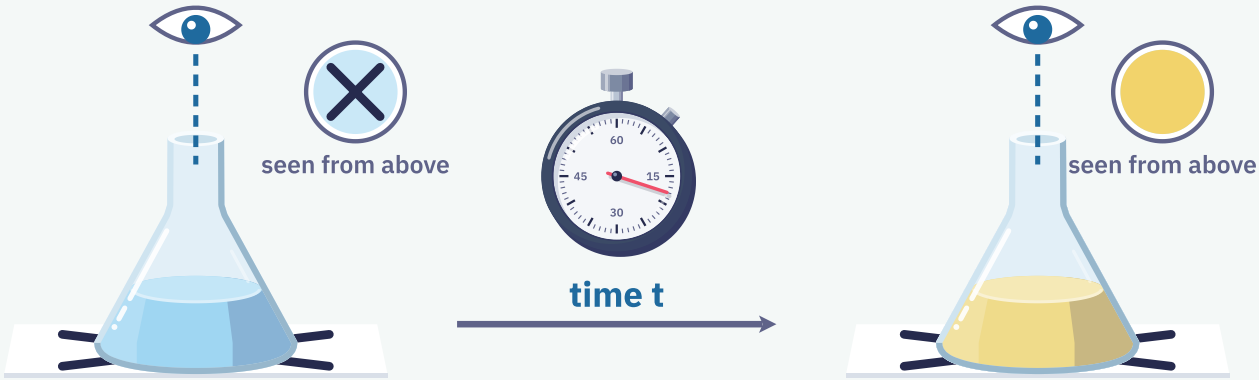
1.2 Reading a rate graph



1. The _____ is the rate. It is steepest at the start, as the reactants are most concentrated.
2. The gradient falls as the reactants are used up, so the reaction _____.
3. The graph _____ when the limiting reactant is used up and the reaction stops.

1.3 Other ways to follow a reaction





- A graph of the **mass of the reaction mixture** against time slopes _____, as the mass of reactants falls. A graph of the mass of a product slopes up.
- When a precipitate forms, time how long it takes for a cross under the flask to **disappear**. A shorter time means a faster rate:

$$\text{rate} ____ \frac{1}{\text{time taken}}$$

$Na_2S_2O_3(aq) + 2HCl(aq) \rightarrow 2NaCl(aq) + H_2O(l) + SO_2(g) + S(s)$

- The yellow precipitate of **sulfur** slowly hides the cross.

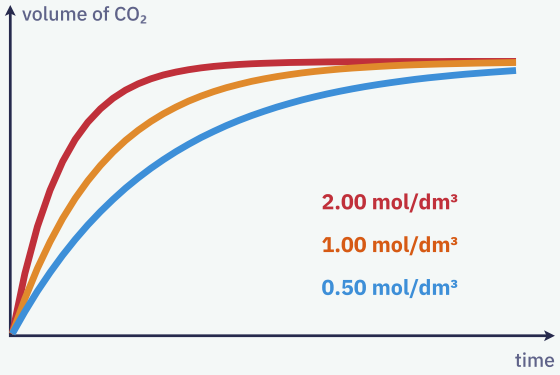
1.4 Planning an investigation

1. Change _____ variable only, e.g. the concentration of the acid.
2. Keep every other variable the _____: the volume of acid, the mass and size of the solid, and the temperature.
3. Measure the volume of gas (or the mass) at _____ with a stopwatch, until it stops changing.
4. Plot the results for each experiment on the **same axes** and compare the initial gradients.

EXAM TIP In a plan, name the apparatus (gas syringe or balance, and stopwatch), say what is kept the same, and say how the result shows the rate: a steeper initial gradient, or a shorter time, means a faster rate.

1 Measuring the Rate (continued)

1.5 Comparing rate graphs



- The curve with the _____ _____ shows the fastest rate.
- The **final volume** depends only on the amount of the _____. Here the marble chips are limiting, so all three end at the same volume.
- Halving the amount of the limiting reactant halves the final volume.

WORKED EXAMPLE The gradient of a mass against time graph is positive when the mass of a product is measured, but negative when the mass of the reaction flask and its contents is measured. Why?

Products form as the reaction goes on, so the mass of a product **increases**: the gradient is positive.

The gas given off escapes from the flask, so the mass of the flask and its contents **decreases**: the gradient is negative.

2 What Affects the Rate

2.1 Effective collisions

Particles react only when they collide with energy _____ the **activation energy**, E_a. Such a collision is an _____.



less than the activation energy: they bounce apart



at least the activation energy: they react

- The rate increases when the **frequency of effective collisions** increases.

2 What Affects the Rate (continued)

2.2 Concentration and pressure

- A higher **concentration** means more reacting particles _____. They collide more often, so the frequency of effective collisions increases.
- For **gases**, a higher **pressure** pushes the particles closer together, with the same effect. Pressure has almost no effect on solids and liquids.

Concentration (or pressure, for a gas) ↑

1. Number of reacting particles per _____ ↑
2. → frequency of collisions ↑, and hence frequency of _____ ↑ → rate ↑

2.3 Particle size

- Smaller pieces have a larger _____ for the same mass, so more particles are exposed to collide.

Particle size of the solid ↓

1. Larger total _____ of contact
2. → frequency of effective collisions ↑ → rate ↑

WORKED EXAMPLE Explain, in terms of collisions, why powdered marble reacts faster with dilute hydrochloric acid than the same mass of marble lumps.

Powdered marble has a larger **total surface area** of contact with the acid.

This increases the **frequency of effective collisions**, so the rate is faster.

2.4 Temperature

- Particles gain **kinetic energy** and move faster, so they collide _____.
- More particles have energy _____, so a larger fraction of collisions are effective.
- Both increase the frequency of effective collisions, so the rate increases.

Temperature ↑

1. Particles gain kinetic energy and move faster → frequency of collisions ↑
2. → more particles have energy _____, so a larger fraction of collisions are effective
3. → frequency of effective collisions ↑ → rate ↑

EXAM TIP Explain a rate in three links: what changes about the particles, so they collide more frequently (or with more energy), so the **frequency of effective collisions** increases. Say "frequency", not just "more collisions".

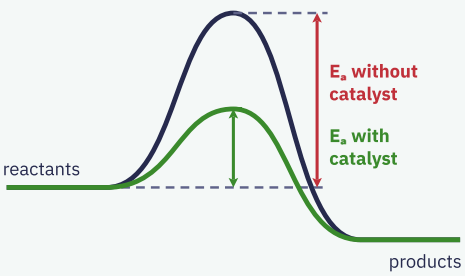
3 Catalysts

3.1 What a catalyst does

DEFINITION A **catalyst** increases the rate of a reaction _____ at the end of the reaction.

- It is not used up, so a small amount works for a long time. Its physical appearance may change.
- It does _____ change the yield of a reaction.
- It is selective: each catalyst speeds up only certain reactions.

3.2 A lower activation energy



- A catalyst provides an _____ with a _____.
- More particles now have energy greater than or equal to the activation energy, so the frequency of effective collisions increases.
- The energies of the reactants and the products are **unchanged**.

Catalyst present

1. Provides an alternative pathway with a lower activation energy
2. → more particles have energy ≥ E_a
3. → frequency of effective collisions ↑ → rate ↑

WORKED EXAMPLE Reaction A has a higher activation energy than reaction B. Under the same conditions, which reaction is faster? Explain your answer.

Reaction **B** is faster.

Its activation energy is lower, so more particles have energy greater than or equal to it.

The **frequency of effective collisions** is higher, so the rate is faster.

3 Catalysts (continued)

3.3 Catalysts in industry and in living things

Catalyst	Used in
_____	the Haber process, to make ammonia
aluminium oxide or silicon dioxide	the cracking of hydrocarbons
_____	making margarine from vegetable oils
platinum, palladium and rhodium	catalytic converters in cars

- Industrial catalysts are usually **transition metals** or their compounds. They let reactions run at lower temperatures and pressures, which saves energy.
- **Enzymes** are _____. They are proteins, and each works best at a particular temperature (often body temperature, 35 to 40 °C) and pH. High temperatures _____ them.