LEARNERS' ALLEY
Designed by Khow Willard

5 Acids, Bases and Salts

SINGAPORE JUNIOR CHEMISTRY OLYMPIAD



Lemon juice, vinegar and stomach acid are acids. Oven cleaner and soap are alkalis. This chapter explains what makes a substance an acid or a base, how acids and bases react, how to measure acidity, how to make salts, and how a titration measures exactly how much acid or alkali a solution holds.

Learning objectives

By the end of this chapter, you can:

Foundation O LEVEL

- describe acids and alkalis in terms of H^+ and OH^- ions, and their reactions
- explain the difference between a strong and a concentrated acid
- use the pH scale and indicators
- classify oxides as acidic, basic, amphoteric or neutral
- use the solubility rules to choose a method of making a salt
- calculate concentrations from titration results

Core A LEVEL

- apply the Arrhenius, Brønsted–Lowry and Lewis theories
- identify conjugate acid–base pairs
- choose an indicator for a titration
- use back-titration to analyse an insoluble solid

Prerequisites

Chapter 1 (filtration and crystallisation), Chapter 3 (ions, formulae, lone pairs and dative bonds) and Chapter 4 (ionic equations, moles and concentration). Fill in the blanks as the lesson goes.

5.1 Acids

DEFINITION An **acid** is a substance that produces _____, when dissolved in water.

When an acid dissolves, its molecules ionise: $HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$. Common acids are hydrochloric acid, HCl, nitric acid, HNO_3, sulfuric acid, H_2SO_4, and ethanoic acid, CH_3COOH, the acid in vinegar. The H^+ ions give all acids the same properties: they taste sour, turn blue litmus red, and their solutions conduct electricity because they contain ions.

5.1.1 The three reactions of acids

Learn these three reactions. Each forms a **salt**, named after the acid: sulfates from sulfuric acid, nitrates from nitric acid and chlorides from hydrochloric acid.

1 acid + metal → salt + hydrogen

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

- **Test**: hydrogen puts out a lighted splint with a 'pop'.
- Only reactive metals react. Copper and silver do not react with dilute acids.
- Lead seems not to react with hydrochloric or sulfuric acid, because a layer of insoluble lead(II) chloride or lead(II) sulfate coats it.

2 acid + base → salt + water

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

- This is **neutralisation** (Section 5.3).

3 acid + carbonate → salt + water + carbon dioxide

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

- **Test**: carbon dioxide turns limewater milky.

WORKED EXAMPLE Write the ionic equation for $CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$.

1. $CaCO_3(s) + 2H^+(aq) + 2Cl^-(aq) \rightarrow Ca^{2+}(aq) + 2Cl^-(aq) + H_2O(l) + CO_2(g)$ — *Split the aqueous ionic compounds. The solid stays whole.*

2. _____ — *Cl^- is a spectator ion.*

Check your understanding

1. Name the salt and the gas formed when zinc reacts with dilute hydrochloric acid.

2. Why does solid citric acid not conduct electricity, while its solution in water does?

5.2 Strong and weak acids

The **strength** of an acid is how far it ionises in water (Figure 5.1).

DEFINITION A **strong acid** is _____ ionised in aqueous solution. A **weak acid** is only _____ ionised.

Hydrochloric, nitric and sulfuric acids are strong: $HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$. Ethanoic acid and citric acid are weak. Only some ethanoic acid molecules ionise, and the ions can join again, so the equation has a reversible arrow: $CH_3COOH(aq) \rightleftharpoons CH_3COO^-(aq) + H^+(aq)$.

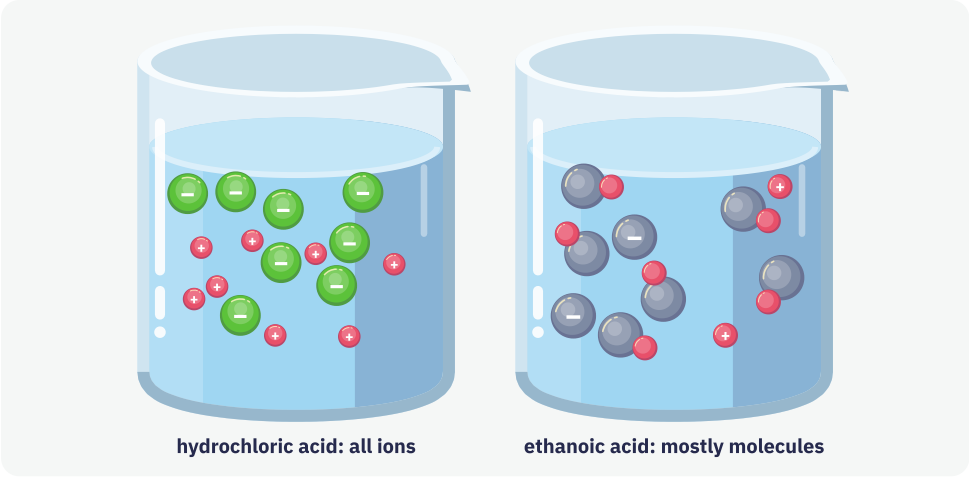


Figure 5.1 A strong acid is fully ionised. A weak acid is mostly un-ionised molecules (numbers for illustration).

EXAM TIP Strong is not the same as concentrated. **Concentration** is how much acid is dissolved in a given volume. A strong acid can be dilute, and a weak acid can be concentrated.

Check your understanding

1. Solutions of hydrochloric acid and ethanoic acid have the same concentration. Which has the higher concentration of H^+ ions, and why?

5.3 Bases and alkalis

DEFINITION A **base** is a metal oxide or metal hydroxide. An **alkali** is a base that is _____ in water, producing _____.

5.3.1 The hydroxide ion

A hydroxide ion, OH^-, is an oxygen atom covalently bonded to a hydrogen atom, with one extra electron. The extra electron gives it a charge of 1−, and its oxygen keeps three lone pairs. Being negative and having lone pairs to spare, OH^- strongly attracts H^+ ions and joins with them to form water:

$$OH^-(aq) + H^+(aq) \rightarrow H_2O(l)$$

This is what makes a hydroxide a base: it removes the H^+ ions that make a solution acidic. The oxide ion, O^{2-}, in a metal oxide does the same, taking two H^+ ions to form one water molecule.

Most bases, such as copper(II) oxide and magnesium hydroxide, are insoluble. Sodium hydroxide, potassium hydroxide and calcium hydroxide dissolve, so they are alkalis. Aqueous ammonia is also an alkali, although it contains no metal: $NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$. Like acids, alkalis are strong or weak. Sodium hydroxide is fully ionised, so it is a strong alkali. Ammonia is only partly ionised, so it is a weak alkali. Alkalis feel soapy and turn red litmus blue.

5.3.2 The two reactions of bases

1 base + acid → salt + water neutralisation

For example, hydrochloric acid is neutralised by aqueous sodium hydroxide:

$$NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$$

Whatever the acid and alkali, the only ions that react are the H^+ of the acid and the OH^- of the alkali (Figure 5.2):

$$\mathbf{H^+(aq) + OH^-(aq) \rightarrow H_2O(l)}$$



Figure 5.2 Hydrochloric acid neutralised by sodium hydroxide: only H^+ and OH^- react.

- Na^+ and Cl^- are spectator ions. They stay in solution as sodium chloride, the salt.
- The salt of a strong acid and a strong alkali, such as sodium chloride or sodium nitrate, gives a **neutral** solution.

2 alkali + ammonium salt → salt + water + ammonia

$$KOH(aq) + NH_4Cl(s) \rightarrow KCl(aq) + H_2O(l) + NH_3(g)$$

- The mixture is **warmed**.
- The OH^- takes an H^+ from the NH_4^+ ion, leaving NH_3.
- **Test**: ammonia has a sharp smell and turns damp red litmus paper blue.

Check your understanding

1. Write the balanced equation for magnesium hydroxide reacting with hydrochloric acid.
2. Is every base an alkali? Explain.

5.4 The pH scale and indicators

The **pH scale**, usually from 0 to 14, shows how acidic or alkaline a solution is (Figure 5.3).

- pH < 7: more H^+ than OH^-, so the solution is **acidic**. The lower the pH, the more acidic.
- pH = 7: equal H^+ and OH^-, so the solution is **neutral**, like pure water.
- pH > 7: more OH^- than H^+, so the solution is **alkaline**.

Each step of one pH unit is a _____ change in the concentration of H^+. A solution of pH 4 has 1000 times more H^+ than one of pH 7.

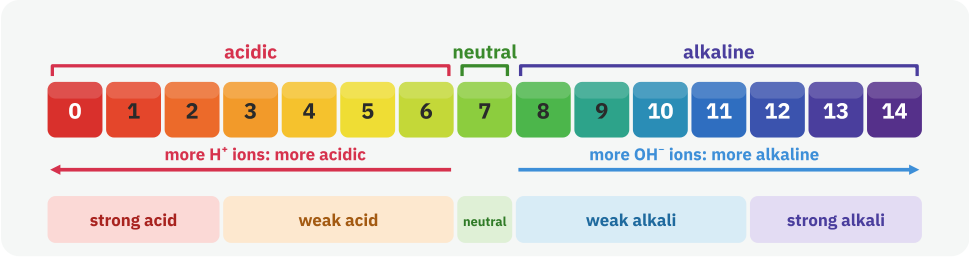


Figure 5.3 The pH scale and the colours of Universal Indicator.

pH can be measured with a pH meter or with an **indicator**, a dye whose colour depends on pH. **Universal Indicator** is a mixture of dyes that gives a different colour at each pH. Other indicators have one colour change over a narrow range:

indicator	colour in acid	changes over pH	colour in alkali
methyl orange	red	3 to 5	yellow
litmus	red	5 to 8	blue
phenolphthalein	colourless	8 to 10	pink
thymolphthalein	colourless	9 to 10.5	blue

Most plants grow best in neutral or slightly acidic soil. Soil that is too acidic is treated with a base such as calcium hydroxide (slaked lime), which neutralises the excess acid.

Check your understanding

1. Solution P has pH 2 and solution Q has pH 5. How many times greater is the concentration of H^+ in P?
2. What colour is methyl orange in a solution of pH 9?

5.5 Oxides

How an oxide behaves depends on whether its element is a metal or a non-metal (Figure 5.4).

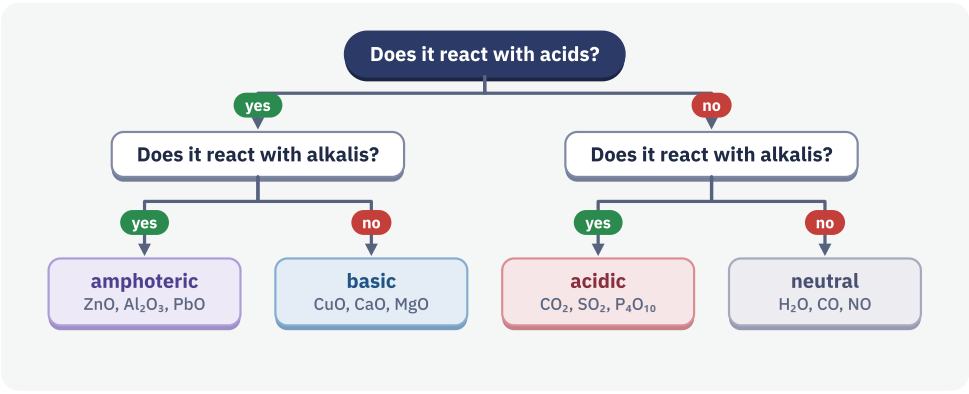


Figure 5.4 Classifying oxides.

1 Basic oxides — oxides of most metals

They react with acids to form a salt and water.

$$MgO(s) + 2HNO_3(aq) \rightarrow Mg(NO_3)_2(aq) + H_2O(l)$$

- Most are insoluble in water.
- Sodium oxide and potassium oxide dissolve to form alkalis: $Na_2O(s) + H_2O(l) \rightarrow 2NaOH(aq)$

2 Amphoteric oxides ZnO, Al_2O_3, PbO

They react with _____ acids and alkalis to form a salt and water.

with an acid: $ZnO(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2O(l)$

with an alkali: $ZnO(s) + 2NaOH(aq) \rightarrow Na_2ZnO_2(aq) + H_2O(l)$

- With sodium hydroxide, zinc oxide forms the salt sodium zincate.

3 Acidic oxides oxides of most non-metals

They react with alkalis to form a salt and water.

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

Most dissolve in water to form acids.

$$SO_2(g) + H_2O(l) \rightarrow H_2SO_3(aq), \text{ sulfurous acid}$$

- SO_3 dissolves to give sulfuric acid, and CO_2 gives carbonic acid.
- Silicon dioxide is insoluble, but it reacts with hot concentrated sodium hydroxide, so it is acidic too.

4 Neutral oxides H_2O, CO, NO

They react with neither acids nor alkalis.

Check your understanding

1. A white oxide is insoluble in water but dissolves in both dilute hydrochloric acid and aqueous sodium hydroxide. What type of oxide is it? Name one possible oxide.

5.6 Theories of acids and bases

CORE · A LEVEL

The definitions so far are the **Arrhenius** theory: an acid dissociates in water to give $H^+(aq)$, and a base gives $OH^-(aq)$. It works only in water, and it cannot explain why ammonia, with no OH group, is a base. Two wider theories fix this.

Brønsted–Lowry theory: an acid is a _____ and a base is a _____.

A proton is far too small to exist on its own. When an acid dissolves, it gives its proton to a lone pair on a water molecule, forming the **hydronium ion**, H_3O^+ (Figure 5.5). $H^+(aq)$ and $H_3O^+(aq)$ mean the same thing. A Brønsted–Lowry base must have a lone pair to accept the proton.

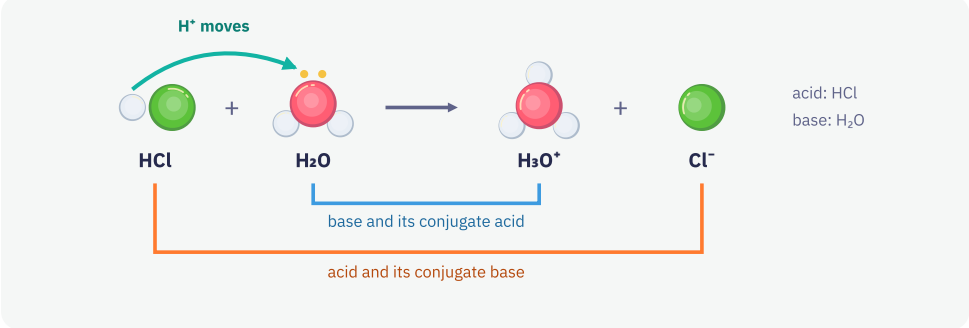


Figure 5.5 Proton transfer from HCl to water, and the two conjugate pairs.

Two species that differ by one H^+ are a **conjugate acid–base pair**. When an acid gives away its proton, what is left is its **conjugate base**. When a base accepts a proton, it becomes its **conjugate acid**. In $HCl + H_2O \rightarrow H_3O^+ + Cl^-$, Cl^- is the conjugate base of HCl, and H_3O^+ is the conjugate acid of H_2O.

WORKED EXAMPLE Identify the conjugate pairs in $NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$.

1. NH_3 gains H^+ to become NH_4^+: base and conjugate acid. *The base accepts the proton.*

2. H_2O loses H^+ to become OH^-: acid and conjugate base. *Here water acts as an acid. With HCl it acted as a base.*

Water can act as an acid or a base, depending on what it reacts with. An acid that can donate one proton per molecule, such as HCl or HNO_3, is **monobasic**. H_2SO_4 is dibasic and H_3PO_4 is tribasic.

Lewis theory: an acid is an _____ and a base is an _____.

Ammonia and boron trifluoride show why this happens (Figure 5.6). Boron is in Group 13, so in BF_3 it shares its three valence electrons and ends up with only _____ electrons round it, two short of an octet. The nitrogen in NH_3 has a lone pair that is not used in bonding. Nitrogen gives this pair to boron, forming a dative bond (Chapter 3), and boron reaches its octet: $NH_3 + BF_3 \rightarrow H_3N{\rightarrow}BF_3$.

- NH_3 donates the electron pair, so it is the **Lewis base**.
- BF_3 accepts the electron pair, so it is the **Lewis acid**.

No proton moves, so this is an acid–base reaction only in the Lewis sense. Every Brønsted–Lowry base is also a Lewis base, because it accepts a proton with a lone pair.

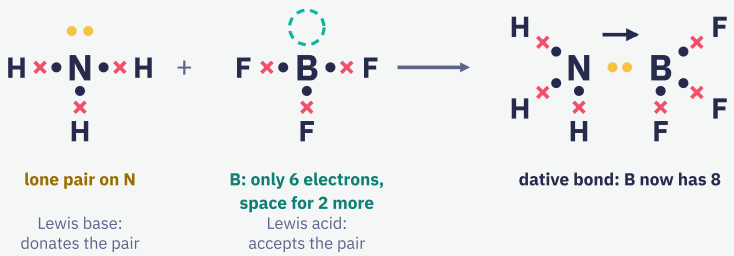


Figure 5.6 A Lewis acid–base reaction. Boron has six outer electrons until nitrogen donates its lone pair. The lone pairs on fluorine are not shown.

Check your understanding

1. Give the conjugate base of H_2SO_4 and the conjugate acid of NH_3.
2. In $HCl(g) + NH_3(g) \rightarrow NH_4Cl(s)$, which reactant is the Brønsted–Lowry acid? Why is this not an acid–base reaction in the Arrhenius sense?

5.7 Salts

A **salt** is an ionic compound made of a cation and an anion. It forms when the H^+ of an acid is replaced by a metal ion or an ammonium ion.

So far, every salt has come from an acid. There is a second way. When two solutions are mixed and the cation from one and the anion from the other form an _____ salt, the salt appears as a solid in the liquid, a **precipitate** (Figure 5.7). This is a **precipitation** reaction:

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

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

Na^+ and NO_3^- are spectator ions and stay in solution. The precipitate is separated by filtration (Chapter 1).

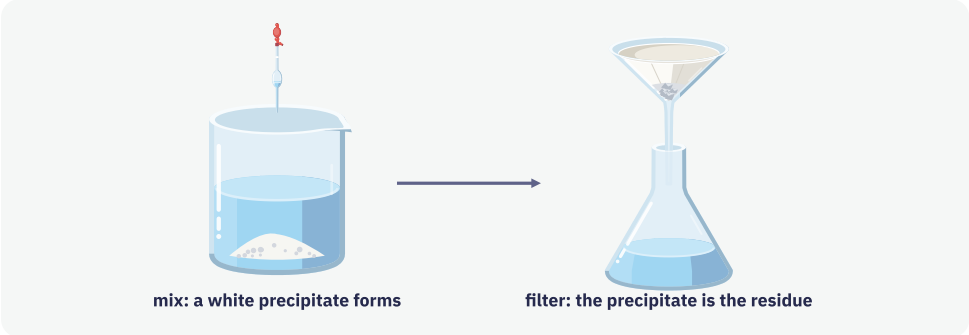


Figure 5.7 Precipitation of barium sulfate.

To choose how to make a salt, first decide whether it dissolves in water.

salts	soluble	insoluble
sodium, potassium and ammonium salts	all	none
nitrates	all	none
chlorides	most	AgCl, $PbCl_2$
sulfates	most	$BaSO_4$, $PbSO_4$, $CaSO_4$ (sparingly soluble)
carbonates	Na_2CO_3, K_2CO_3, $(NH_4)_2CO_3$	all others

There are three methods (Figure 5.8).

- **Acid + excess insoluble solid**, for a soluble salt: add a metal, insoluble base or insoluble carbonate to the acid until no more reacts. The excess solid makes sure all the acid is used, and it is then filtered off. Heat the filtrate until it is saturated, cool it to crystallise the salt, filter, wash the crystals with a little cold water and dry them. Very reactive metals are too dangerous, and unreactive metals do not react.
- **Titration**, for a soluble salt from soluble starting materials: sodium, potassium and ammonium salts. Neither reactant can be filtered off, so the exact volumes are found by titration with an indicator (Section 5.8). The reaction is repeated with the same volumes and no indicator, and the salt is crystallised.
- **Precipitation**, for an insoluble salt: mix a solution containing its cation with a solution containing its anion. $Ba(NO_3)_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2NaNO_3(aq)$. Filter, wash the precipitate with distilled water and dry it.

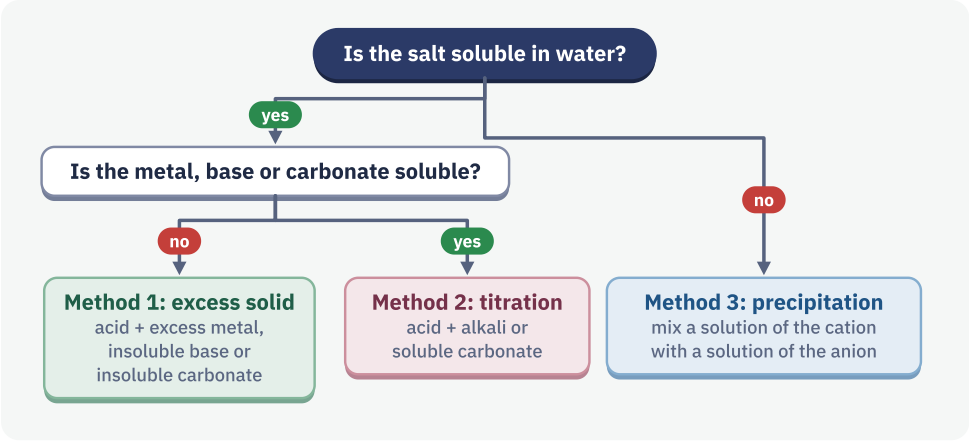


Figure 5.8 Choosing a method to make a salt.

WORKED EXAMPLE Describe how to make lead(II) chloride from lead(II) carbonate, dilute nitric acid and aqueous potassium chloride.

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

$PbCl_2$ is insoluble, so precipitate it. First make a soluble lead salt.

2. Add excess KCl(aq) to the filtrate: $Pb(NO_3)_2(aq) + 2KCl(aq) \rightarrow PbCl_2(s) + 2KNO_3(aq)$. Filter, wash and dry.

Every nitrate is soluble.

Check your understanding

1. Suggest starting materials and a method to make (a) copper(II) sulfate, (b) potassium nitrate and (c) silver chloride.

5.8 Acid–base titrations

A titration measures the volume of one solution that reacts exactly with a fixed volume of another (Figure 5.9). A pipette delivers a fixed volume of one solution into a conical flask with a few drops of indicator. The other solution is added from a burette until the indicator just changes colour.

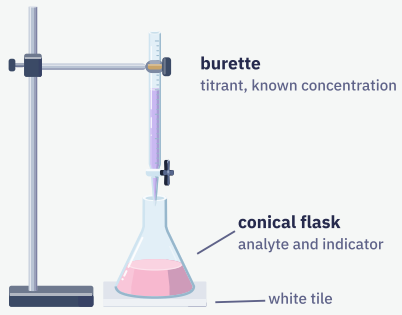


Figure 5.9 A titration.

- A **standard solution** is one whose concentration is accurately known.
- The **equivalence point** is reached when the reactants have reacted exactly in the ratio of the equation. The **end-point** is when the indicator changes colour. A good indicator changes colour at the equivalence point.

WORKED EXAMPLE 5.00 g of impure sodium hydroxide is dissolved and made up to 1.00 dm^3. 25.0 cm^3 of this solution is neutralised by 15.15 cm^3 of 0.102 mol dm^{-3} sulfuric acid. Find the percentage purity of the sodium hydroxide.

1. $2NaOH + H_2SO_4 \rightarrow Na_2SO_4 + 2H_2O$	*Ratio NaOH : H_2SO_4 = 2 : 1.*
2. $n_{H_2SO_4} = 0.102 \times 0.01515 = 1.545 \times 10^{-3}$ mol, so n_{NaOH} in 25.0 $cm^3 = 3.091 \times 10^{-3}$ mol	*Volumes in dm^3.*
3. In 1.00 dm^3: $3.091 \times 10^{-3} \times 1000/25.0$ = 0.1236 mol, mass = 0.1236 × 40.0 = 4.945 g	*Scale up from 25.0 cm^3 to the whole solution.*
4. Purity = 4.945 ÷ 5.00 × 100% = _____	*Pure mass ÷ sample mass.*

Some salts crystallise with water molecules held in their lattice, called **water of crystallisation**. Copper(II) sulfate crystals are $CuSO_4 \cdot 5H_2O$. A titration can find the number of water molecules.

WORKED EXAMPLE 8.65 g of $Na_2CO_3 \cdot nH_2O$ is made up to 250 cm^3. 25.0 cm^3 of this solution reacts with 24.00 cm^3 of 0.125 mol dm^{-3} sulfuric acid: $Na_2CO_3 + H_2SO_4 \rightarrow Na_2SO_4 + CO_2 + H_2O$. Find n.

1. $n_{H_2SO_4} = 0.125 \times 0.02400 = 3.000 \times 10^{-3}$ mol = $n_{Na_2CO_3}$ in 25.0 cm^3	*Ratio 1 : 1.*
2. In 250 cm^3: 0.03000 mol, so M_r = 8.65 ÷ 0.03000 = 288.3	*One $Na_2CO_3 \cdot nH_2O$ per Na_2CO_3.*
3. 106.0 + 18.0n = 288.3, so n = 10.1, which is _____	*n is a whole number.*

5.8.1 Choosing an indicator CORE · A LEVEL

At the equivalence point the pH changes very quickly. The indicator must change colour within that sudden change, which depends on whether the acid and base are strong or weak.

titration	indicator
strong acid and strong base	methyl orange, phenolphthalein or thymolphthalein
strong acid and weak base	methyl orange
weak acid and strong base	phenolphthalein or thymolphthalein
weak acid and weak base	no suitable indicator

5.8.2 Back-titration CORE · A LEVEL

An insoluble solid, such as calcium carbonate or magnesium hydroxide, cannot be titrated directly. Instead, a known excess of acid is added so that all the solid reacts. The acid left over is then titrated with a standard alkali. The acid that reacted with the solid is found by difference (Figure 5.10).

amount of acid that reacted = amount added − amount left over

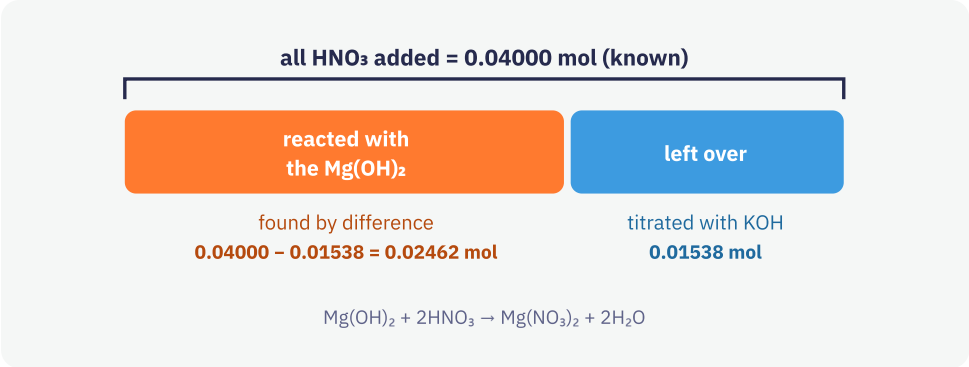


Figure 5.10 Back-titration: the acid added splits into the part used by the solid and the part left over.

WORKED EXAMPLE An antacid tablet of mass 1.00 g contains insoluble $Mg(OH)_2$. It is added to 100 cm^3 of 0.400 mol dm^{-3} HNO_3 (excess). The leftover acid needs 20.50 cm^3 of 0.750 mol dm^{-3} KOH. Find the percentage of $Mg(OH)_2$ in the tablet.

1. $n_{HNO_3\ added} = 0.400 \times 0.100 = 0.04000$ mol. $n_{HNO_3\ left\ over} = n_{KOH} = 0.750 \times 0.02050 = 0.01538$ mol — *KOH reacts 1 : 1 with HNO_3.*
2. $n_{HNO_3\ reacted} = 0.04000 - 0.01538 = 0.02462$ mol — *By difference.*
3. $Mg(OH)_2 + 2HNO_3 \rightarrow Mg(NO_3)_2 + 2H_2O$, so $n_{Mg(OH)_2} = 0.01231$ mol — *Ratio 1 : 2.*
4. Mass = $0.01231 \times 58.3 = 0.7177$ g, so _____ — *M_r of $Mg(OH)_2$ = 58.3.*

Check your understanding

1. 25.0 cm^3 of aqueous sodium hydroxide is neutralised by 15.00 cm^3 of 0.250 mol dm^{-3} hydrochloric acid. Find the concentration of the sodium hydroxide.
2. Which indicator would you use to titrate aqueous ammonia with hydrochloric acid? Why?
3. 2.50 g of impure $CaCO_3$ is added to 50.0 cm^3 of 1.00 mol dm^{-3} HCl (excess). The leftover acid needs 20.0 cm^3 of 0.500 mol dm^{-3} NaOH. Find the percentage of $CaCO_3$ (M_r 100.1).

Answers to Check your understanding

5.1

1. Zinc chloride and hydrogen: $Zn + 2HCl \rightarrow ZnCl_2 + H_2$
2. In the solid there are no free ions. In water the acid ionises, and the ions carry the current.

5.2

1. Hydrochloric acid. It is fully ionised, while most ethanoic acid molecules stay un-ionised.

5.3

1. $Mg(OH)_2(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + 2H_2O(l)$
2. No. Only bases that dissolve in water are alkalis. Copper(II) oxide is a base but not an alkali.

5.4

1. Three pH units, so 10 × 10 × 10 = 1000 times.
2. Yellow

5.5

1. Amphoteric, for example zinc oxide or aluminium oxide.

5.6

1. HSO_4^- and NH_4^+
2. HCl, because it gives a proton to NH_3. There is no water and no $H^+(aq)$ or $OH^-(aq)$, so the Arrhenius theory does not apply.

5.7

1. (a) Excess copper(II) oxide or carbonate with dilute sulfuric acid, filter, crystallise. (b) Potassium hydroxide or carbonate with dilute nitric acid by titration, then crystallise. (c) Precipitation: silver nitrate solution with a soluble chloride such as sodium chloride, filter, wash, dry.

5.8

1. $n_{HCl} = 0.250 \times 0.01500 = 3.750 \times 10^{-3}$ mol $= n_{NaOH}$. $c = 3.750 \times 10^{-3} \div 0.0250 = 0.150$ mol dm^{-3}
2. Methyl orange. Hydrochloric acid is strong and ammonia is a weak base, so the sudden pH change at the equivalence point lies in the acidic range.
3. HCl added 0.05000 mol, left over 0.01000 mol, so 0.04000 mol reacted. $n_{CaCO_3} = 0.02000$ mol, mass = 2.002 g, purity = 80.1%