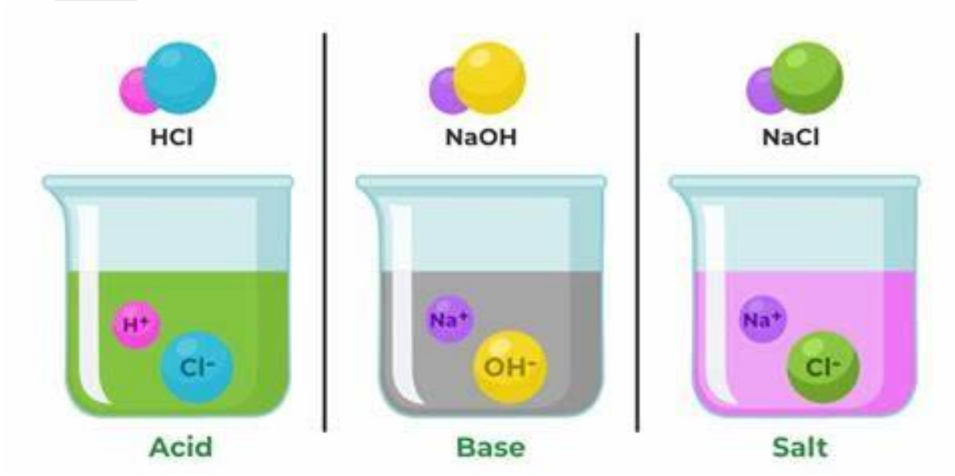
Cambridge OL

Chemistry

CODE: (5070)

Chapter 08

Acids, bases and salts

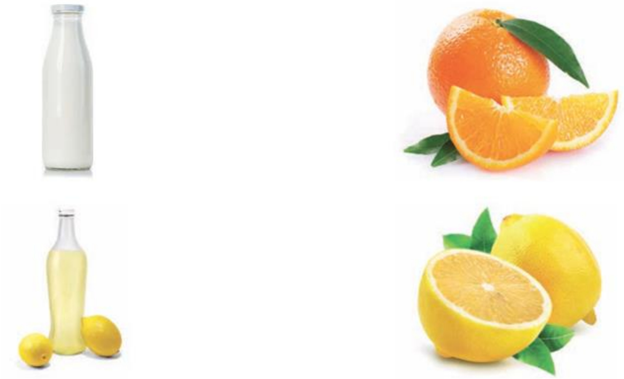


8.1 Acids and alkalis

Acids are ubiquitous in our environment, and their properties make them recognizable as such.

Some common alkaline substances are shown in Figure 8.2.

Chemists use indicators, dyes extracted from natural sources like litmus, to determine acidity in liquids. These **indicators** change color when added to acids or alkalis, making it dangerous to taste.



▲ **Figure 8.1** What do all these foods and drinks have in common?

▼ **Table 8.1** Indicators and their colours in acid and alkaline solution

Indicator	Colour in acid solution	Colour in alkaline solution
Blue litmus	Red	Blue
Methyl orange	Pink	Yellow
Thymolphthalein	Colourless	Blue
Red litmus	Red	Blue

▲ **Figure 8.2** Some common alkaline substances

The pH scale, developed by Søren Sørenson, is a color-coded indicator used to determine a substance's acidity or alkalineness. It ranges from 0 to 14, with a pH below 7 indicating acidity, a pH greater than 7 indicating alkalineness, and a pH of 7 indicating neutrality, such as water.

▲ **Figure 8.3** Indicators tell you if a substance is acid or alkaline

The pH of a substance can be measured using a universal indicator or a pH meter, which indicates the concentration of hydrogen ions in the solution, based on the redder color.

▲ **Figure 8.5** A digital pH meter

The Brønsted–Lowry theory

This theory defined:

» An **acid** as an H+ ion (or proton) donor

» A **base** as an H+ ion (or proton) acceptor

Key definitions

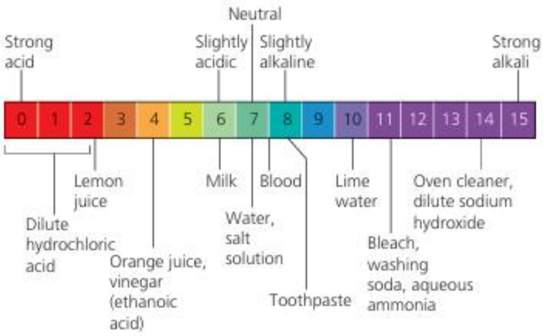
Acids are proton donors.

Bases are proton acceptors.

So, the water, in the aqueous acid solution, is behaving as a Brønsted–Lowry base and accepting an H+ ion. Generally:

$$\underset{\text{acid}}{HA(aq)} + \underset{\text{base}}{H_2O(l)} \rightarrow H_3O^+(aq) + A^-(aq)$$

If a substance can behave both as a proton acceptor (base) and a proton donor (acid) then it is called **amphoteric.**



a The pH scale

b Universal indicator in solution, showing the colour range

▲ **Figure 8.4**

The relative strengths of acids and bases

The relative strength of an acid and base is determined by comparing their dissociation in water, while the strength of a base is determined by its dissociation in water.

Strong and weak acids

A typical strong acid is hydrochloric acid. It is formed by dissolving hydrogen chloride gas in water. In hydrochloric acid, the ions formed separate completely.

$$\text{hydrogen chloride} \xrightarrow{\text{water}} \underset{\text{ions}}{\text{hydrogen}} + \underset{\text{ions}}{\text{chloride}}$$

$$HCl(g) \xrightarrow{\text{water}} H^+(aq) + Cl^-(aq)$$

Key definition

A **strong acid** is an acid that is completely dissociated in aqueous solution. Hydrochloric acid is a strong acid: $HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$

A **strong acid**, such as hydrochloric acid, sulfuric acid, and nitric acid, breaks down hydrogen chloride molecules to form H+ and Cl– ions.

One mole of sulfuric acid completely dissociates to give 2 moles of hydrogen ions and 1 mole of sulfate ions:

$$H_2SO_4(aq) \rightarrow 2H^+(aq) + SO_4^{2-}(aq)$$

When strong acids are neutralised by strong alkalis, the following reaction takes place between hydrogen ions and hydroxide ions.

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

A **weak acid**, like ethanoic acid in vinegar, produces few hydrogen ions and has a higher pH than a strong acid, but less than 7.

$$\text{ethanoic acid} \overset{\text{water}}{\rightleftharpoons} \text{hydrogen ions} + \text{ethanoate ions}$$

$$CH_3COOH(aq) \overset{\text{water}}{\rightleftharpoons} H^+(aq) + CH_3COO^-(aq)$$

Key definition

A **weak acid** is an acid that is partially dissociated in aqueous solution. Ethanoic acid is a weak acid: $CH_3COOH(aq) \rightleftharpoons CH_3COO^-(aq) + H^+(aq)$

Acids, such as ethanoic acid, are weak acids that produce hydrogen ions when in aqueous solution. They are reversible and have fewer hydrogen ions than strong acids. Examples of weak acids include citric acid, carbonic acid, sulfurous acid, and ascorbic acid. Solutions of weak acids are poorer conductors of electricity and have slower reactions with metals, bases, and metal carbonates. Strong acids are still strong even in dilute solutions, while weak acids remain weak.

▲ **Figure 8.6** Sulfurous acid is found in acid rain. It is a weak acid and is oxidised to sulfuric acid (a strong acid). Acid rain damages the environment quite badly

Neutralisation reactions

Indigestion, caused by stomach acid, can be treated with indigestion remedies containing a substance that neutralizes the acid. In laboratories, a common acid like hydrochloric acid can be neutralised using sodium hydroxide. When sodium hydroxide solution is added to the acid, the pH increases. When equal volumes of hydrochloric acid and sodium hydroxide are added, a neutral solution forms, indicating the acid has undergone a **neutralisation reaction**.

$$\text{hydrochloric acid} + \text{sodium hydroxide} \rightarrow \text{sodium chloride} + \text{water}$$

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

Key definition

A **neutralisation reaction** occurs between an acid and a base to produce water; $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$.

As we have shown, when both hydrochloric acid and sodium hydroxide dissolve in water, the ions separate completely. We may therefore write:

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

They are called spectator ions and are usually taken out of the equation. The equation now becomes:

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

An **ionic equation** describes the reaction between acids and alkali in aqueous solutions, resulting in the formation of neutral water, where H+ ions cause acidity and OH- ions cause alkalinity.

8.2 Formation of salts

Salts can be classified as those which are soluble in water or those which are insoluble in water. The general solubility rules for salts are:

- All sodium, potassium and ammonium salts are soluble
- All nitrates are soluble
- All chlorides are soluble, except lead and silver
- All sulfates are soluble, except barium, calcium and lead
- All carbonates and hydroxides are insoluble, except sodium, potassium and ammonium.

▼ **Table 8.2** Types of salt and the acids they are formed from

Acid	Type of salt	Example
Carbonic acid	Carbonates	Sodium carbonate (Na_2CO_3)
Ethanoic acid	Ethanoates	Sodium ethanoate (CH_3COONa)
Hydrochloric acid	Chlorides	Potassium chloride (KCl)
Nitric acid	Nitrates	Potassium nitrate (KNO_3)
Sulfuric acid	Sulfates	Sodium sulfate (Na_2SO_4)

Types of oxides

Amphoteric oxides are rare, acidic oxides that react with acids and alkalis to produce salts. Examples include zinc oxide and aluminium oxide. Non-metal oxides produce hydrogen ions in aqueous solutions, while metal oxides are bases.

8.3 Methods of preparing soluble salts

There are four general methods of preparing soluble salts. These involve the reaction of an acid with:

- Excess metal
- Excess insoluble carbonate
- Excess insoluble base
- An alkali by titration.

Acid + metal

$$\text{acid} + \text{metal} \longrightarrow \text{salt} + \text{hydrogen}$$

MAZIT metals, including magnesium, aluminium, zinc, iron, and tin, are commonly used in salt preparation methods. An experimental method involves adding excess magnesium ribbon to dilute hydrochloric acid, ensuring all reactants are left, and producing hydrogen gas through effervescence. This method is dangerous when using reactive metals like sodium.

$$\begin{array}{ccccc} \text{hydrogen ions} & + & \text{electrons} & \longrightarrow & \text{hydrogen gas} \\ & & \text{(from metal)} & & \\ 2H^+ & + & 2e^- & \longrightarrow & H_2(g) \end{array}$$

$$\begin{array}{ccccccc} \text{magnesium} & + & \text{hydrochloric} & \longrightarrow & \text{magnesium} & + & \text{hydrogen} \\ & & \text{acid} & & \text{chloride} & & \\ Mg(s) & + & 2HCl(aq) & \longrightarrow & MgCl_2(aq) & + & H_2(g) \end{array}$$

The magnesium chloride solution is evaporated slowly to form a **saturated** solution of the salt (Figure 8.8).

▲ **Figure 8.7** The excess magnesium is filtered in this way

The magnesium chloride solution is tested by dipping a cold glass rod, forming salt crystals at the end, which are then filtered and dried between clean tissues.

Acid + carbonate

$$\text{acid} + \text{carbonate} \longrightarrow \text{salt} + \text{water} + \text{carbon dioxide}$$

This method can be applied to any metal carbonate and acid, producing soluble salt, and effervescence is observed when excess copper (II) carbonate is added to nitric acid.

$$\begin{array}{ccccccccc} \text{copper(II)} & + & \text{nitric} & \longrightarrow & \text{copper(II)} & + & \text{carbon} & + & \text{water} \\ \text{carbonate} & & \text{acid} & & \text{nitrate} & & \text{dioxide} & & \\ CuCO_3(s) & + & 2HNO_3(aq) & \longrightarrow & Cu(NO_3)_2(aq) & + & CO_2(g) & + & H_2O(l) \end{array}$$

Metal carbonates contain carbonate ions, CO3 2−. In this reaction the carbonate ions react with the hydrogen ions in the acid.

$$\begin{array}{ccccccc} \text{carbonate} & + & \text{hydrogen} & \longrightarrow & \text{carbon} & + & \text{water} \\ \text{ions} & & \text{ions} & & \text{dioxide} & & \\ CO_3^{2-}(aq) & + & 2H^+(aq) & \longrightarrow & CO_2(g) & + & H_2O(l) \end{array}$$

▲ **Figure 8.8** The solution of magnesium nitrate is concentrated by slow evaporation

Acid + alkali (soluble base)

acid + alkali → salt + water

Titration is a method used to prepare salts of reactive metals like potassium or sodium, by using an alkali containing the metal. Most metal oxides and hydroxides are **insoluble bases**, while some are **alkalis**, or **soluble bases**, which dissolve in water to produce OH-(aq) ions. Insoluble bases, on the other hand, are those that do not dissolve in water. This indirect method is used when adding the metal directly to the acid.

Key definitions
Alkalis are soluble **bases**.
Aqueous solutions of alkalis contain OH^- ions.

▼ **Table 8.3** Examples of soluble and insoluble bases

Soluble bases (alkalis)	**Insoluble bases**
Sodium hydroxide (NaOH)	Iron(III) oxide (Fe_2O_3)
Potassium hydroxide (KOH)	Copper(II) oxide (CuO)
Calcium hydroxide ($Ca(OH)_2$)	Lead(II) oxide (PbO)
Ammonia solution ($NH_3(aq)$)	Magnesium oxide (MgO)

Bases neutralize acids, producing salt and water. They are soluble (alkali) or insoluble (insoluble). They accept hydrogen ions and are found in metal oxides, hydrogenates, and ammonia solutions. Salts can only be formed if the base is soluble.

A neutralisation reaction involves adding acid to alkali using a **burette** until an indicator changes color, called the **endpoint**. Once neutralized, the same volume of acid is added to the alkali without the indicator.

▲ **Figure 8.9** The acid is added to the alkali until the indicator just changes colour

Acid + insoluble base

acid + base → salt + water

This method prepares a salt of unreactive metals like lead or copper, neutralizing the acid using a specific metal oxide when direct metal-acid reactions are not feasible.

The method involves neutralizing sulfuric acid with copper (II) oxide, resulting in copper (II) sulfate, similar to the process for metal carbonate and acid, but with potential warming of reactants.

▲ **Figure 8.10** Citric acid has been used on the right-hand side of this piece of copper metal to remove the oxide coating on its surface. The acid reacts with the oxide coating to form water, leaving shiny copper metal

▲ **Figure 8.11** After slow evaporation to concentrate the solution, the solution is left to crystallise. Crystals of copper(II) sulfate are produced

sulfuric acid + copper(II) oxide → copper(II) sulfate + water

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

Metal oxides contain the oxide ion, O^{2-}. The ionic equation for this reaction is therefore:

$$2H^+(aq) + O^{2-}(s) \rightarrow H_2O(l)$$

or

$$CuO(s) + 2H^+(aq) \rightarrow Cu^{2+}(aq) + H_2O(l)$$

Reaction of a base with ammonium salts

Ammonia gas, NH_3, can be produced by heating ammonium salt with a base like calcium hydroxide in a laboratory setting.

$$\begin{array}{ccccccccc} \text{calcium hydroxide} & + & \text{ammonium chloride} & \rightarrow & \text{calcium chloride} & + & \text{water} & + & \text{ammonia} \\ Ca(OH)_2(s) & + & 2NH_4Cl & \rightarrow & CaCl_2 & + & H_2O & + & 2NH_3 \end{array}$$

The ammonia produced can be detected as being formed by its pungent odour and by turning damp red litmus blue

8.4 Preparing insoluble salts

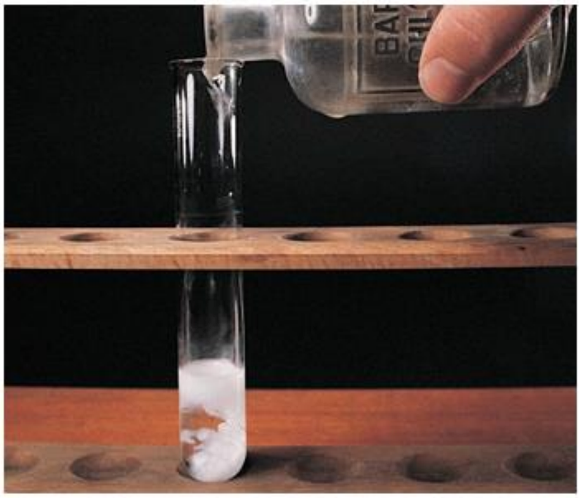
The methods mentioned can create soluble water salts, while insoluble salts like barium sulfate can be prepared through precipitation using barium chloride and sodium sulfate. The precipitate can be filtered, washed, and dried.

$$\begin{array}{ccccccc} \text{barium chloride} & + & \text{sodium sulfate} & \rightarrow & \text{barium sulfate} & + & \text{sodium chloride} \end{array}$$

$$BaCl_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2NaCl(aq)$$

The ionic equation for this reaction is:

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



▲ **Figure 8.12** When barium chloride solution is added to sodium sulfate, a white precipitate of barium sulfate forms

This method may be summarised as follows:

$$\begin{array}{ccccccc} \text{soluble salt} & + & \text{soluble salt} & \rightarrow & \text{insoluble salt} & + & \text{soluble salt} \\ (AX) & + & (BY) & \rightarrow & (BX) & + & (AY) \end{array}$$

8.5 Testing for different salts

Sometimes we want to analyse a salt and find out what is in it. There are simple chemical tests which allow us to identify the anion part of the salt.

Testing for a sulfate ($SO_4{}^{2-}$)

Barium sulfate, an insoluble salt, forms a white precipitate when a suspected sulfate solution is mixed with a soluble barium salt like barium chloride.

$$\begin{array}{ccccc} \text{barium ion} & + & \text{sulfate ion} & \rightarrow & \text{barium sulfate} \\ Ba^{2+}(aq) & + & SO_4^{2-}(aq) & \rightarrow & BaSO_4(s) \end{array}$$

Add dilute hydrochloric acid to a mixture, and if precipitate doesn't dissolve, it's barium sulfate, while if it dissolves, it may be a sulfite containing $SO_3{}^{2-}$ ion.

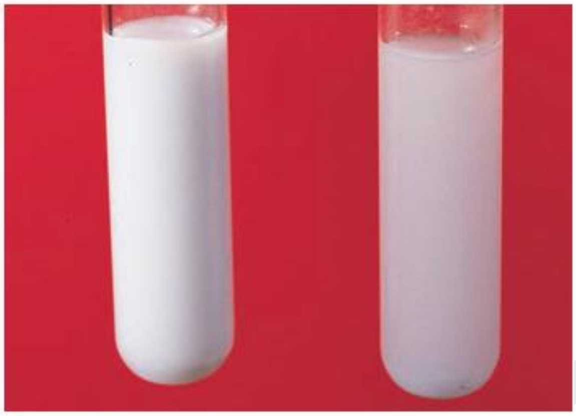
Testing for a chloride (Cl−), a bromide (Br−) or an iodide (I−)

Silver chloride is an insoluble salt, so adding dilute nitric acid to a solution of suspected chloride creates an aqueous acidic solution, followed by a soluble silver salt solution.

$$\text{chloride ion} + \text{silver ion} \rightarrow \text{silver chloride}$$

$$Cl^-(aq) + Ag^+(aq) \rightarrow AgCl(s)$$

An alternative test for iodide ions is the addition of lead nitrate solution to the iodide which results in a bright yellow precipitate of lead iodide, PbI_2



▲ **Figure 8.13** If left to stand the white precipitate of silver chloride goes grey. This photochemical change plays an essential part in black and white photography

▲ **Figure 8.14** AgCl, a white precipitate, AgBr, a cream precipitate, and AgI, a yellow precipitate

Testing for a carbonate

If a small amount of an acid is added to some of the suspected carbonate (either solid or in solution) then effervescence occurs. If it is a carbonate then carbon dioxide gas is produced, which will turn limewater 'milky' (a cloudy white precipitate of calcium carbonate forms, see Chapter 14, p. 227)

$$\begin{array}{ccccccc} \text{carbonate} & + & \text{hydrogen} & \rightarrow & \text{carbon} & + & \text{water} \\ \text{ions} & & \text{ions} & & \text{dioxide} & & \\ CO_3^{2-}(aq) & + & 2H^+(aq) & \rightarrow & CO_2(g) & + & H_2O(l) \end{array}$$

Testing for a nitrate

By adding aqueous sodium hydroxide and then aluminium foil and warming gently, nitrates are reduced to ammonia. The ammonia can be identified using damp indicator paper, which turns blue

$$\begin{gathered} 3NO_3^-(aq) + 8Al(s) + 5OH^-(aq) + 18H_2O(l) \\ \downarrow \\ 3NH_3(g) + 8[Al(OH)_4]^-(aq) \end{gathered}$$

8.6 Water of crystallisation

Hydrates, a type of salt, are formed when they crystallise from solution. These salts, like sodium chloride, copper carbonate, and sodium nitrate, crystallise in their **anhydrous** forms, but many also produce hydrates when they crystallise from solution. The shape of a crystal hydrate is significantly influenced by the presence of **water of crystallisation**.

▼ **Table 8.4** Examples of crystal hydrates.

Salt hydrate	Formula
Cobalt(II) chloride hexahydrate	$CoCl_2.6H_2O$
Copper(II) sulfate pentahydrate	$CuSO_4.5H_2O$
Iron(II) sulfate heptahydrate	$FeSO_4.7H_2O$
Magnesium sulfate heptahydrate	$MgSO_4.7H_2O$

When many hydrates are heated, the water of crystallisation is driven away. For example, if crystals of copper (II) sulfate hydrate (blue) are heated strongly, they lose their water of crystallisation. Anhydrous copper (II) sulfate remains as a white powder:

copper(II) sulfate pentahydrate → anhydrous copper(II) sulfate + water

$$CuSO_4.5H_2O(s) \rightarrow CuSO_4(s) + 5H_2O(g)$$

When water is added to anhydrous copper (II) sulfate, the reverse process occurs. It turns blue and the pentahydrate is produced (Figure 8.16). This is an extremely exothermic process.

$$CuSO_4(s) + 5H_2O(l) \rightarrow CuSO_4.5H_2O(s)$$

Because the colour change only takes place in the presence of water, the reaction is used to test for the presence of water. These processes give a simple example of a reversible reaction:

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

Titration

Titration is a method used to prepare a soluble salt by reacting an acid with an alkali, which can also determine the concentration of the alkali used.

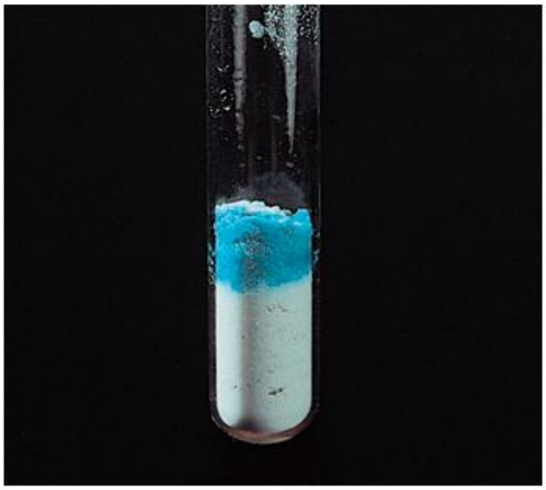
▲ **Figure 8.15** Hydrate crystals (left to right): cobalt nitrate, calcium nitrate and nickel sulfate (top) and manganese sulfate, copper sulfate and chromium potassium sulfate (bottom)

Key definitions

A hydrated substance, or **hydrate** is one that is chemically combined with water.

An **anhydrous** substance is one containing no water.

Water of crystallisation is the water molecules present in crystals, e.g. $CuSO_4.5H_2O$ and $CoCl_2.6H_2O$.



▲ **Figure 8.16** Anhydrous copper(II) sulfate is a white powder which turns blue when water is added to it

Revision questions

21. Nov/2021/Paper_21/No.9

This question is about ammonia and nitrates.

(a) State the source of the hydrogen and nitrogen used in the manufacture of ammonia by the Haber process.

source of hydrogen ....

source of nitrogen ....
[2]

(b) Ammonia is formed when aqueous ammonium sulfate is heated with sodium hydroxide.

$$(NH_4)_2SO_4 + 2NaOH \rightarrow 2NH_3 + Na_2SO_4 + 2H_2O$$

A student adds 4.50 g of sodium hydroxide to 50.0 cm^3 of 1.25 mol / dm^3 aqueous ammonium sulfate.

Show by calculation that the ammonium sulfate is in excess.

(c) When dilute sodium hydroxide is electrolysed, the hydroxide ions are converted to oxygen and water at the anode.

Construct the ionic equation for the reaction taking place at the anode.

.... [1]

(d) The first stage in the test for nitrate ions is to heat aluminium and sodium hydroxide with a solution containing nitrate ions.

State the type of chemical reaction which occurs when a solution containing nitrate ions is heated with aluminium and sodium hydroxide.

.... [1]

(e) Nitrates from fertilisers cause eutrophication when they are leached from soils into rivers.

Describe the process of eutrophication.

22. Nov/2021/Paper_22/No.9

This question is about sulfuric acid and sulfates.

(a) **(i)** Sulfur is one of the raw materials used in the Contact process to make sulfuric acid. Name two other raw materials used to make sulfuric acid.

.... and .... [1]

(ii) Name the catalyst used in the Contact process.

.... [1]

(b) The electrolysis of dilute sulfuric acid using inert electrodes produces oxygen at the anode.

(i) Define the term *electrolysis*.

....

....

.... [2]

(ii) Construct the equation for the reaction taking place at the cathode.

.... [1]

(c) Dilute sulfuric acid reacts with sodium hydroxide.

$$H_2SO_4 + 2NaOH \rightarrow Na_2SO_4 + 2H_2O$$

A student adds 0.76 g of solid sodium hydroxide to 45 cm^3 of 0.20 mol / dm^3 sulfuric acid.

Show by calculation that the sodium hydroxide is in excess.

(d) Aqueous sodium hydroxide is warmed with ammonium sulfate. State the names of the three products formed in this reaction.

9. Nov/2021/Paper_12/No.8a(ii)

(ii) Describe how to prepare pure dry crystals of barium nitrate from aqueous barium nitrate.

10. Nov/2021/Paper_22/No.3b

(b) In the past, ink was made from a mixture containing iron(II) ions and tannic acid.

(i) Describe a test for iron(II) ions.

test ....

observations ....

[2]

(ii) The ink darkens when used on paper. This is because iron(II) ions are oxidised to iron(III) ions.

Write the ionic equation for this reaction.

.... [1]

(iii) After a time, the ink fades because of a hydrolysis reaction which is catalysed by acids.

State how a catalyst increases the rate of a chemical reaction.

....

.... [1]

11. Nov/2021/Paper_22/No.6

Methanoic acid and ethanoic acid are weak acids.

(a) What does the term *weak* mean, when applied to acids?

....

.... [1]

(b) Methanoic acid, HCO_2H, reacts with magnesium powder.

(i) Construct the equation for this reaction.

.... [1]

(ii) State and explain how the rate of this reaction changes when the experiment is repeated using a piece of magnesium ribbon with the same mass as the powder.

All other conditions stay the same.

Include in your answer ideas about collisions between particles.

....

....

.... [2]

(c) Methanoic acid reacts with propanol, C_3H_7OH, to form an ester.

Name and draw the structure of this ester, showing all of the atoms and all of the bonds.

name ....

structure

(d) Ethanoic acid is present in vinegar.

(i) Name the organic compound which is converted to ethanoic acid when vinegar is made.

.... [1]

(ii) What type of chemical reaction is this?

.... [1]

(e) The table shows some properties of four carboxylic acids.

carboxylic acid	formula	density in g/cm^3	boiling point in °C
methanoic acid	HCO_2H	1.22	101
ethanoic acid	CH_3CO_2H		118
propanoic acid	$C_2H_5CO_2H$	0.99	141
butanoic acid	$C_3H_7CO_2H$	0.96	164

(i) Predict the density of ethanoic acid.

.... g/cm^3 [1]

(ii) Describe and explain the change in the boiling point as the number of carbon atoms in a molecule increases.

16. Jun/2020/Paper_21/No.5

Hydrochloric acid, HC*l*, reacts with barium hydroxide, $Ba(OH)_2$, as shown.

$$2HCl(aq) + Ba(OH)_2(aq) \rightarrow BaCl_2(aq) + 2H_2O(l)$$

A sample of 25.0 cm^3 of 0.0500 mol / dm^3 $Ba(OH)_2$ is placed in a beaker.

Dilute HC*l* is added slowly, from a burette, to the $Ba(OH)_2(aq)$ in the beaker.

A pH probe is used to measure the pH of the solution in the beaker until a total of 40.0 cm^3 of dilute HC*l* is added.

The table shows how the pH of the solution in the beaker changes.

volume of dilute HC*l* added / cm^3	pH of the solution in the beaker
0.0	13.0
5.0	12.9
10.0	12.5
15.0	11.6
20.0	7.0
25.0	3.0
30.0	1.6
35.0	1.1
40.0	0.9

(a) Explain, in terms of the ions present, why the pH of the solution in the beaker changes from 13.0 to 0.9.

(b) Use the data in the table to state the volume of dilute HC*l* that just neutralises all of the sample of $Ba(OH)_2(aq)$.

volume of dilute HC*l* .... cm^3 [1]

(c) Use your answer to **(b)** to calculate the concentration, in mol / dm^3, of the dilute HC*l*.