8 Acids, bases and salts

FOCUS POINTS
★ What are acids and bases, and how do they behave?
★ What are the solubility rules?
★ How can we prepare soluble salts?

In this chapter, you will look at acids and bases, their properties and their reactions. You will look at the use of indicators to tell whether a solution is acidic or alkaline, and you will be introduced to indicators other than universal indicator solution, which you have probably already met.

The reactions we will look at in this chapter produce ionic solids called salts. Salts are very important to our lives: for example, ammonium nitrate is the most-used fertiliser on our planet and magnesium sulfate is used in medicines. This chapter will guide you through different methods of making salts and which reactants you would use. For example, how do you make ammonium nitrate? Which reactants will you need and which of the different methods you will read about should you use?

By the end of this chapter you should be able to answer these questions. You will also become more confident in writing chemical equations and using these equations to work out the quantities of materials you need to start with and how much of the product you should make.

8.1 Acids and alkalis

All the substances shown in Figure 8.1 contain an **acid** of one sort or another. Acids are certainly all around us. What properties do these substances have which make you think that they are acids or contain acids?

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

Some common alkaline substances are shown in Figure 8.2.

▲ **Figure 8.2** Some common alkaline substances

It would be too dangerous to taste a liquid to find out if it was acidic. Chemists use substances called **indicators** which change colour when they are added to acids or alkalis. Many indicators are dyes which have been extracted from natural sources, such as litmus.

Methyl orange, thymolphthalein and blue and red litmus are all indicators used in titrations (see p. 128). Table 8.1 shows the colours they turn in acids and alkalis.

▼ **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

These indicators tell chemists whether a substance is acid or alkaline (Figure 8.3). To obtain an idea of how acidic or alkaline a substance is, we use another indicator known as a universal indicator. This indicator is a mixture of many other indicators. The colour shown by this indicator can be matched against a **pH scale**. The pH scale was developed by a Scandinavian chemist called Søren Sørenson. The pH scale runs from below 0 to 14. A substance with a pH of less than 7 is an acid. One with a pH of greater than 7 is alkaline. One with a pH of 7 is said to be neither acid nor alkaline: it is neutral. Water is the most common example of a neutral substance. Figure 8.4 shows the universal indicator colour range along with everyday substances with their particular pH values.

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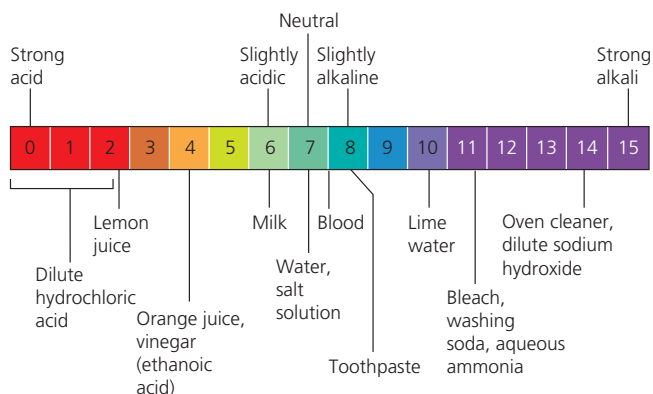
a The pH scale

b Universal indicator in solution, showing the colour range

▲ **Figure 8.4**

Using universal indicator allows us to determine how concentrated one acid solution is compared to another. The redder the colour, the more acidic the solution is. Solutions which are very acidic contain higher concentrations of hydrogen ions ($H^+(aq)$).

Another way in which the pH of a substance can be measured is by using a pH meter (Figure 8.5). The pH electrode is placed into the solution and a pH reading is given on the digital display.

▲ **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.

The theory explains why a pure acid behaves differently from its aqueous solution, since for an acid to behave as an H^+ ion donor it must have another substance present to accept the H^+ ion. So the water, in the aqueous acid solution, is behaving as a Brønsted–Lowry base and accepting an H^+ ion. Generally:

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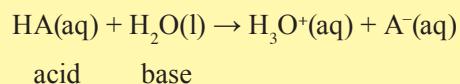
$$\text{acid} \qquad \text{base}$$

If a substance can behave both as a proton acceptor (base) and a proton donor (acid) then it is called **amphoteric**, for example, aluminium oxide, Al_2O_3 and zinc oxide, ZnO.

The relative strengths of acids and bases

The relative strength of an acid is found by comparing one acid with another. The strength of any acid depends upon how many molecules dissociate (or ionise) when the acid is dissolved in water. The relative strength of a base is found by comparing one base with another and is again dependent upon the dissociation of the base in aqueous solution.
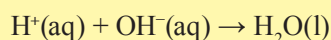
Strong and weak acids

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.

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

ions   ions

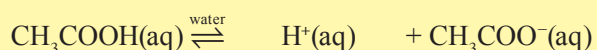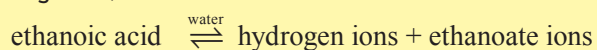$HCl(g) \qquad \xrightarrow{\text{water}} \quad H^+(aq) \; + \; Cl^-(aq)$
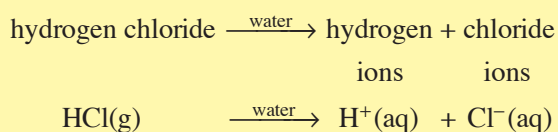
> **Key definition**
>
> A **strong acid** is an acid that is completely dissociated in aqueous solution. Hydrochloric acid is a strong acid: $HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$.

For hydrochloric acid, *all* the hydrogen chloride molecules break up to form H^+ ions and Cl^- ions. Any acid that behaves in this way is termed a **strong acid**. Both sulfuric acid and nitric acid also behave in this way and are therefore also termed strong acids. All these acids have a high concentration of hydrogen ions in solution ($H^+(aq)$) and have a low pH. Their solutions are good conductors of electricity and they react quickly with metals, bases and metal carbonates.

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,** such as ethanoic acid, which is found in vinegar, produces few hydrogen ions when it dissolves in water compared with a strong acid of the same concentration. It is only partially dissociated. Its solution has a higher pH than a strong acid, but still less than 7.

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

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

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

The $\rightleftharpoons$ sign means that the reaction is reversible. This means that if the ethanoic acid molecule breaks down to give hydrogen ions and ethanoate ions then they will react together to re-form the ethanoic acid molecule. The fact that fewer ethanoic acid molecules dissociate compared with a strong acid, and that the reaction is reversible, means that few hydrogen ions are present in the solution. Other examples of weak acids are citric acid, found in oranges and lemons; carbonic acid, found in soft drinks; sulfurous acid (acid rain) (Figure 8.6) and ascorbic acid (vitamin C).

▲ **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

Solutions of weak acids are poorer conductors of electricity and have slower reactions with metals, bases and metal carbonates.

All acids when in aqueous solution produce hydrogen ions, $H^+(aq)$. To say an acid is a strong acid does not mean it is concentrated. The *strength* of an acid tells you how easily it dissociates (ionises) to produce hydrogen ions. The *concentration* of an acid indicates the proportions of water and acid present in aqueous solution. It is important to emphasise that a strong acid is still a strong acid even when it is in dilute solution and a weak acid is still a weak acid even if it is concentrated.

Neutralisation reactions

A common situation involving neutralisation of an acid is when you suffer from indigestion. This is caused by acid in your stomach irritating the stomach lining or throat. Normally it is treated by taking an indigestion remedy containing a substance which will react with and neutralise the acid.

In the laboratory, if you wish to neutralise a common acid such as hydrochloric acid you can use an alkali such as sodium hydroxide. If the pH of the acid is measured when some sodium hydroxide solution is added to it, the pH increases. If equal volumes of the same concentration of hydrochloric acid and sodium hydroxide are added to one another, the resulting solution is found to have a pH of 7. The acid has undergone a **neutralisation reaction** and a neutral solution has been formed.

| hydrochloric acid | + | sodium hydroxide | → | sodium chloride | + | water |

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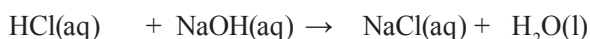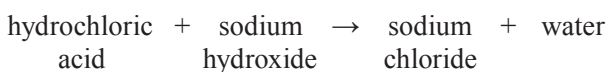
> **Key definition**
>
> A **neutralisation reaction** occurs between an acid and a base to produce water; $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$.

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

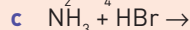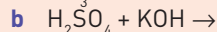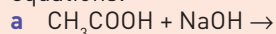$$H^+(aq)Cl^-(aq) + Na^+(aq)OH^-(aq) \rightarrow Na^+(aq) + Cl^-(aq) + H_2O(l)$$
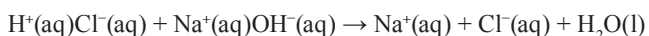
You will notice that certain ions are unchanged on either side of the equation. They are called spectator ions and are usually taken out of the equation. The equation now becomes:

$$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$$
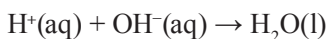
This type of equation is known as an **ionic equation**. The reaction between any acid and any alkali in aqueous solution produces water and can be summarised by this ionic equation. It shows the ion which causes acidity ($H^+(aq)$) reacting with the ion which causes alkalinity ($OH^-(aq)$) to produce neutral water ($H_2O(l)$). Aqueous solutions of acids contain H^+ ions and aqueous solutions of alkalis contain OH^- ions.

> **Test yourself**
>
> 1 Complete the following balanced chemical equations.
>   a $CH_3COOH + NaOH \rightarrow$
>   b $H_2SO_4 + KOH \rightarrow$
>   c $NH_3 + HBr \rightarrow$
>   In each case, name the acid and the base.
>
> 2 Explain the terms 'concentrated' and 'strength' as applied to acids.
>
> 3 Explain what part water plays in the acidity of a solution.
> 4 a Write a chemical equation to represent the neutralisation of sulfuric acid by sodium hydroxide.
>   b Reduce this to an ionic equation.
>   c Account for any difference you see between the ionic equation you have written and the one shown on the left for the reaction of hydrochloric acid and sodium hydroxide.

8.2 Formation of salts

In the example on this page, sodium chloride was produced as part of the neutralisation reaction. Compounds formed in this way are known as salts. A salt is a compound that has been formed when all the hydrogen ions of an acid have been replaced by metal ions or by the ammonium ion (NH_4^+).

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.

If the acid being neutralised is hydrochloric acid, salts called chlorides are formed. Other types of salts can be formed with other acids. A summary of the different types of salt along with the acid they are formed from is shown in Table 8.2.

▼ **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

There are three different types of oxides. Non-metal oxides, such as sulfur dioxide, SO_2, and carbon dioxide are acidic. In aqueous solution they produce aqueous hydrogen ions, $H^+(aq)$. Metal oxides, however, are basic oxides. If these oxides are soluble they will dissolve in water to produce aqueous hydroxide ions, $OH^-(aq)$. For example, copper(II) oxide, CuO and calcium oxide, CaO.

The third type of oxides, and the rarest, are those described as **amphoteric oxides**. These oxides react with both acids and alkalis to produce salts. Examples of amphoteric oxides are zinc oxide, ZnO, and aluminium oxide, Al_2O_3.

8.3 Methods of preparing soluble salts

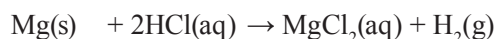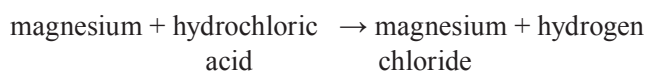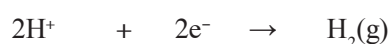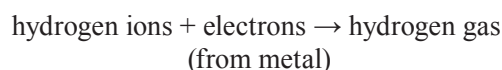There are four general methods of preparing soluble salts. These involve the reaction of an acid with:

» excess metal
» excess insoluble carbonate
» excess insoluble base
» an alkali by titration.

Acid + metal

$$\text{acid} + \text{metal} \rightarrow \text{salt} + \text{hydrogen}$$
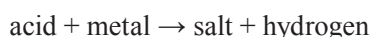
This method can only be used with the less reactive metals such as aluminium. It would be very dangerous to use a reactive metal, such as sodium, in this type of reaction. The metals usually used in this method of salt preparation are the **MAZIT** metals, that is, **m**agnesium, **a**luminium, **z**inc, **i**ron and **t**in. A typical experimental method is given below.

Excess magnesium ribbon is added to dilute hydrochloric acid. By using an excess of magnesium ribbon we are making sure that all of the acid has reacted and some magnesium is left at the end of the reaction. During this addition an effervescence is observed due to the production of hydrogen gas. In this reaction the hydrogen ions from the hydrochloric acid gain electrons from the metal atoms as the reaction proceeds.

$$\text{hydrogen ions} + \text{electrons} \rightarrow \text{hydrogen gas}$$
$$\text{(from metal)}$$

$$2H^+ \quad + \quad 2e^- \quad \rightarrow \quad H_2(g)$$

$$\text{magnesium} + \text{hydrochloric} \rightarrow \text{magnesium} + \text{hydrogen}$$
$$\text{acid} \qquad \qquad \text{chloride}$$

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$

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

The excess magnesium is removed by filtration (Figure 8.7).

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

The hot concentrated magnesium chloride solution produced is tested by dipping a cold glass rod into it. If salt crystals form at the end of the rod the solution is ready to crystallise and is left

to cool. Any crystals produced on cooling are filtered and dried between clean tissues.
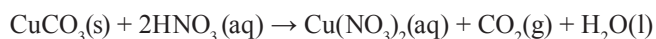

▲ **Figure 8.8** The solution of magnesium nitrate is concentrated by slow evaporation

Acid + carbonate

acid + carbonate → salt + water + carbon dioxide

This method can be used with any metal carbonate and any acid, providing the salt produced is soluble. The typical experimental procedure is similar to that carried out for an acid and a metal. For example, copper(II) carbonate would be added in excess to dilute nitric acid. Effervescence would be observed due to the production of carbon dioxide.

| copper(II) carbonate | + | nitric acid | → | copper(II) nitrate | + | carbon dioxide | + | water |

$$CuCO_3(s) + 2HNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + CO_2(g) + H_2O(l)$$

Metal carbonates contain carbonate ions, CO_3^{2-}. In this reaction the carbonate ions react with the hydrogen ions in the acid.

carbonate ions + hydrogen ions → carbon dioxide + water

$$CO_3^{2-}(aq) + 2H^+(aq) \rightarrow CO_2(g) + H_2O(l)$$
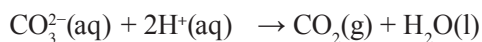
When excess carbonate has been added, you will see it collect at the bottom of the beaker and the effervescence will stop.

The excess copper(II) carbonate can be filtered off and the copper(II) nitrate solution is evaporated slowly to form a saturated solution of the salt.

The hot concentrated copper(II) nitrate solution produced is tested by dipping a cold glass rod into it. If salt crystals form at the end of the rod the solution is ready to crystallise and is left to cool. Any crystals produced on cooling are filtered and dried between clean tissues.

Acid + alkali (soluble base)

acid + alkali → salt + water

Titration is generally used for preparing the salts of very reactive metals, such as potassium or sodium. It would be too dangerous to add the metal directly to the acid. In this case, we solve the problem indirectly and use an alkali which contains the reactive metal whose salt we wish to prepare.

Most metal oxides and hydroxides are insoluble bases: they are said to be basic. A few metal oxides and hydroxides that do dissolve in water to produce $OH^-(aq)$ ions are known as **alkalis**, or **soluble bases**. If the metal oxide or hydroxide does not dissolve in water, it is known as an **insoluble base**.

> **Key definitions**
>
> **Alkalis** are soluble bases. Aqueous solutions of alkalis contain OH⁻ ions.

A **base** is a substance which neutralises an acid, producing a salt and water as the only products. If the base is soluble, the term 'alkali' can be used, but there are several bases which are insoluble. It is also a substance which accepts a hydrogen ion (see p. 118). In general, most metal oxides and hydroxides (as well as ammonia solution) are bases. Some examples of soluble and insoluble bases are shown in Table 8.3. Salts can be formed by this method only if the base is soluble.

▼ **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)

In this neutralisation reaction both reactants are in solution, so a special technique called titration is required. Acid is slowly and carefully added to a measured volume of alkali using a burette (Figure 8.9) until the indicator, for example thymolphthalein, changes colour.

An indicator is used to show when the alkali has been neutralised completely by the acid. This is called the end-point. Once you know where the end-point is, you can add the same volume of acid to the measured volume of alkali but this time without the indicator.


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

For example, the salt sodium chloride can be made by reacting sodium hydroxide with hydrochloric acid using this method.

hydrochloric + sodium → sodium + water
acid hydroxide chloride

$$HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$$
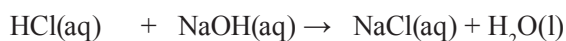
As previously discussed on p. 120, this reaction can best be described by the ionic equation:

$$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$$
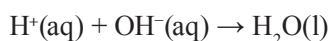
The sodium chloride solution which is produced can then be evaporated slowly to obtain the salt.

The hot concentrated sodium chloride solution produced is tested by dipping a cold glass rod into it. If salt crystals form at the end of the rod, the solution is ready to crystallise and is left to cool. Any crystals produced on cooling are filtered and dried between clean tissues.

Acid + insoluble base

acid + base → salt + water
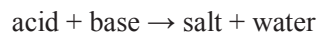
This method can be used to prepare a salt of an unreactive metal, such as lead or copper. In these cases it is not possible to use a direct reaction of the metal with an acid so the acid is neutralised using the particular metal oxide (Figure 8.10).


▲ **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

The method is generally the same as that for a metal carbonate and an acid, though some warming of the reactants may be necessary. An example of such a reaction is the neutralisation of sulfuric acid by copper(II) oxide to produce copper(II) sulfate (Figure 8.11).
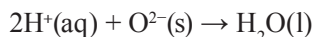

▲ **Figure 8.11** After slow evaporation to concentrate the solution, the solution is left to crystallise. Crystals of copper(II) sulfate are produced

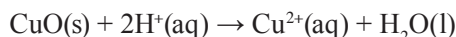sulfuric + copper(II) → copper(II) + water
acid oxide sulfate

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

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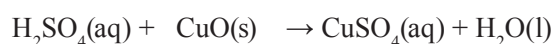
Metal oxides contain the oxide ion, O^{2-}. The ionic equation for this reaction is therefore:

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

or

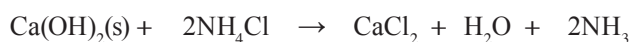$$CuO(s) + 2H^+(aq) \rightarrow Cu^{2+}(aq) + H_2O(l)$$

Reaction of a base with ammonium salts

Small quantities of ammonia gas, NH_3, can be produced by heating any ammonium salt, such as ammonium chloride, with a base, such as calcium hydroxide. This reaction can be used to make ammonia in a laboratory.

calcium + ammonium → calcium + water + ammonia
hydroxide   chloride   chloride

$$Ca(OH)_2(s) + 2NH_4Cl \rightarrow CaCl_2 + H_2O + 2NH_3$$

The ammonia produced can be detected as being formed by its pungent odour and by turning damp red litmus blue.
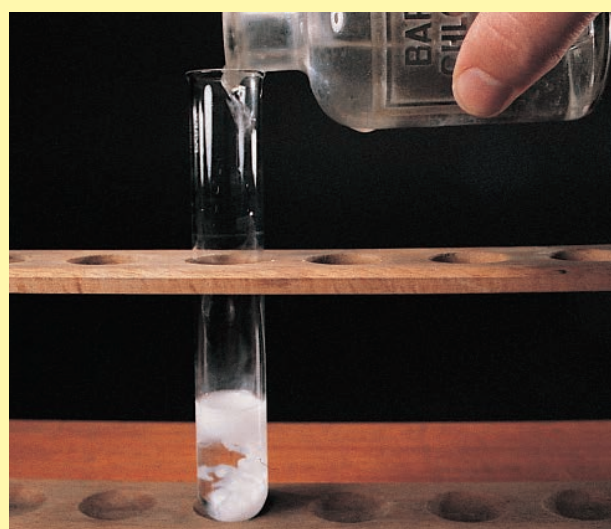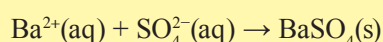
8.4 Preparing insoluble salts

The methods described above can be used to make a salt that is soluble in water. If a salt that is insoluble in water needs to be prepared, a different technique is needed.

An insoluble salt, such as barium sulfate, can be made by precipitation. In this case, solutions of the two chosen soluble salts are mixed (Figure 8.12). To produce barium sulfate, barium chloride and sodium sulfate can be used. The barium sulfate precipitate can be filtered off, washed with distilled water and dried. The reaction that has occurred is:

barium + sodium → barium + sodium
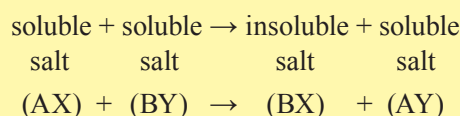chloride   sulfate   sulfate   chloride

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

The ionic equation for this reaction is:
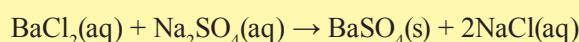
$$Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$$

▲ **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:

soluble + soluble → insoluble + soluble
salt   salt   salt   salt
(AX) + (BY) →   (BX)   + (AY)

It should be noted that even salts like barium sulfate dissolve to a very small extent. For example, 1 dm^3 of water will dissolve 2.2×10^{-3} g of barium sulfate at 25°C. This substance and substances like it are said to be sparingly soluble.

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-})

You have seen that barium sulfate is an insoluble salt (p. 120). Therefore, if you take a solution of a suspected sulfate and add it to a solution of a soluble barium salt (such as barium chloride) then a white precipitate of barium sulfate will be produced.

barium ion + sulfate ion → barium sulfate

$$Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$$
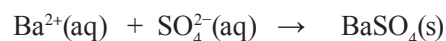
A few drops of dilute hydrochloric acid are also added to this mixture. If the precipitate does not dissolve, then it is barium sulfate and the unknown salt was in fact a sulfate. If the precipitate does dissolve, then the unknown salt may have been a sulfite (containing the SO_3^{2-} ion).

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

Earlier in this chapter you saw that silver chloride is an insoluble salt (p. 120). Therefore, if you take a solution of a suspected chloride and add to it a small volume of dilute nitric acid, to make an aqueous acidic solution, followed by a small amount of a solution of a soluble silver salt (such as silver nitrate), a white precipitate of silver chloride will be produced.

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

$$Cl^-(aq) \quad + \quad Ag^+(aq) \quad \rightarrow \quad AgCl(s)$$
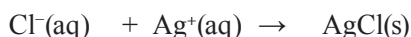
If left to stand, the precipitate goes grey (Figure 8.13).

▲ **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

In a similar way, a bromide and an iodide will react to produce either a cream precipitate of silver bromide (AgBr) or a yellow precipitate of silver iodide (AgI) (Figure 8.14).

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

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.

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. 236).

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

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.

$$3NO_3^-(aq) + 8Al(s) + 5OH^-(aq) + 18H_2O(l)$$
$$\downarrow$$
$$3NH_3(g) + 8[Al(OH)_4]^-(aq)$$
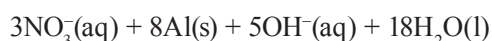
In the reaction the nitrate ion is reduced, as oxygen is removed from the nitrogen atom, and it gains hydrogen to form ammonia, NH_3. The gain of hydrogen is also a definition of reduction.

8.6 Water of crystallisation

Some salts, such as sodium chloride, copper carbonate and sodium nitrate, crystallise in their **anhydrous** forms (without water). However, many salts produce **hydrates** when they crystallise from solution. A hydrate is a salt which incorporates water into its crystal structure. This water is referred to as **water of crystallisation**. The shape of the crystal hydrate is very much dependent on the presence of water of crystallisation. Some examples of crystal hydrates are given in Table 8.4 and shown in Figure 8.15.

Key definitions

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

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

▼ **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	$MgSO_4.7H_2O$

▲ **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)

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 → anhydrous copper(II) + water
pentahydrate   sulfate

$$CuSO_4.5H_2O(s) \rightarrow CuSO_4(s) + 5H_2O(g)$$
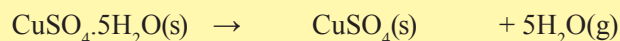
When water is added to anhydrous copper(II) sulfate, the reverse process occurs. It turns blue and the pentahydrate is produced (Figure 8.16). This is an extremely exothermic process.

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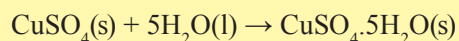
Key definition

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
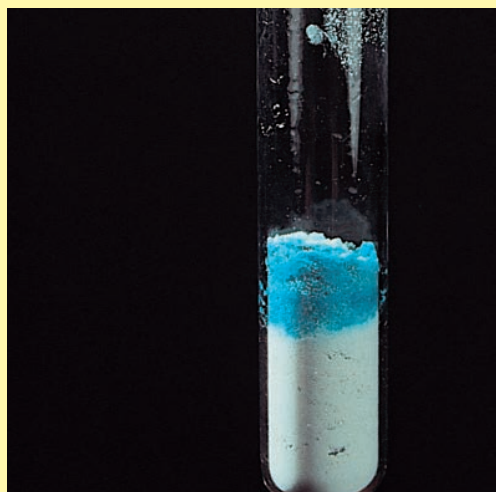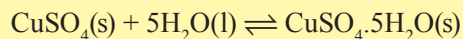
Because the colour change only takes place in the presence of water, the reaction is used to test for the presence of water.

These processes give a simple example of a reversible reaction:

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

Going further

Calculation of water of crystallisation

Sometimes it is necessary to work out the percentage, by mass, of water of crystallisation in a hydrated salt.

? Worked example

Calculate the percentage by mass of water in the salt hydrate $MgSO_4.7H_2O$. (A_r: H = 1; O = 16; Mg = 24; S = 32)

M_r for $MgSO_4.7H_2O$

$$= 24 + 32 + (4 \times 16) + (7 \times 18) = 246$$

The mass of water as a fraction of the total mass of hydrate

$$= \frac{126}{246}$$

The percentage of water present

$$= \frac{126}{246} \times 100$$

$$= 51.2\%$$

Test yourself

8 Calculate the percentage by mass of water in the following salt hydrates:
   a $CuSO_4.5H_2O$ b $Na_2CO_3.10H_2O$
   c $Na_2S_2O_3.5H_2O$.

   (A_r: H = 1; O = 16; Na = 23; S = 32; Cu = 63.5)

Practical skills

Preparation of potassium nitrate

For safe experiments/demonstrations which are related to this chapter, please refer to the *Cambridge IGCSE Chemistry Practical Skills Workbook*, which is also part of this series.

Safety

● Eye protection must be worn.
● Take care using indicator solution as it can stain skin and clothing.

The apparatus below can be used to prepare potassium nitrate crystals.

Safety filler
Burette
Stand
Pipette
Conical flask
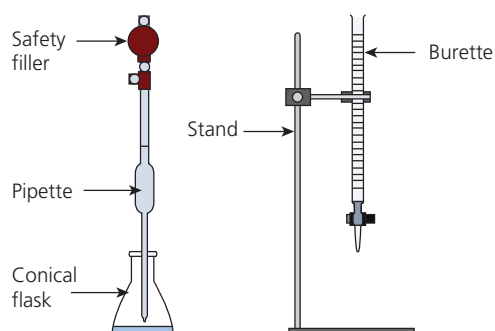
1 Clamp a burette vertically in a clamp stand and fill it up with some dilute nitric acid solution. Allow some of the acid to run through the jet of the burette. Record the initial burette reading.

2 Using a pipette and safety filler, place 25 cm³ of the potassium hydroxide solution into a 250 cm³ conical flask.

3 Put 4 drops of the indicator thymolphthalein into the potassium hydroxide in the flask. The solution will turn blue.

4 Slowly add the dilute nitric acid from the burette, swirling the flask all the time, until the blue colour disappears. Record the final burette reading and hence the amount of dilute nitric acid you needed to add to neutralise the potassium hydroxide.

5 Pour some of the solution formed in step 4 from the flask into an evaporating basin and heat until half of the solution has evaporated.

6 Set aside and allow the crystals to form slowly.

Now answer the following questions:

1 Why is it important to clamp the burette vertically?

2 Why is it important to fill up the jet of the burette before taking the initial burette reading?

3 Why is it important to use a safety filler when using the pipette?

4 How do you take the initial volume reading on the burette?

5 Why is the swirling of the flask important?

6 How could you tell that the solution has become saturated as it is evaporated?

7 In step 5, why is it important to not evaporate all the water away?

8 How can the crystals be dried?

Titration

On p. 122 you saw that it was possible to prepare a soluble salt by reacting an acid with a soluble base (alkali). The method used was that of titration. Titration can also be used to find the concentration of the alkali used. In the laboratory, the titration of hydrochloric acid with sodium hydroxide is carried out in the following way.

1 25 cm³ of sodium hydroxide solution is pipetted into a conical flask to which a few drops of thymolphthalein indicator have been added (Figure 8.17). Thymolphthalein is blue in alkaline conditions but colourless in acid.
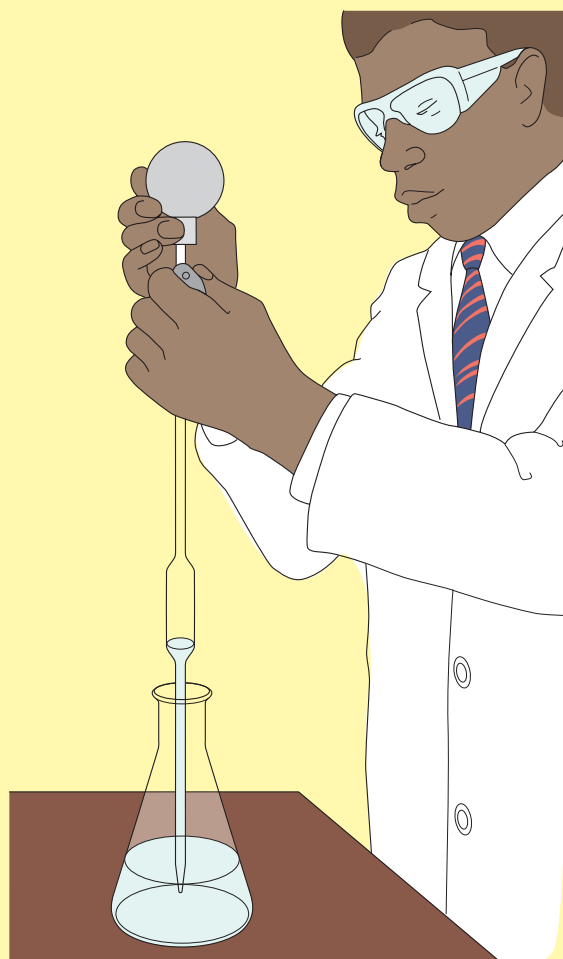
▲ **Figure 8.17** Exactly 25.0 cm³ of sodium hydroxide solution is pipetted into a conical flask

2 A 0.10 mol/dm^3 solution of hydrochloric acid is placed in the burette using a filter funnel until it is filled up exactly to the zero mark (Figure 8.18).

Dilute hydrochloric acid

Burette

Note: Remove filter funnel before titrating
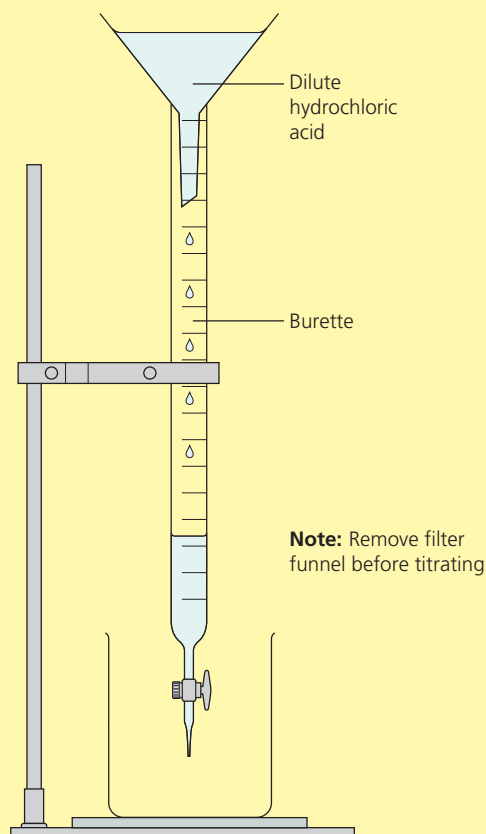
▲ **Figure 8.18** The burette is filled up to the zero mark with a 0.10 mol/dm^3 solution of hydrochloric acid

3 The filter funnel is now removed.
4 The hydrochloric acid is added to the sodium hydroxide solution in small quantities – usually no more than 0.5 cm^3 at a time (Figure 8.19). The contents of the flask must be swirled after each addition of acid for thorough mixing.
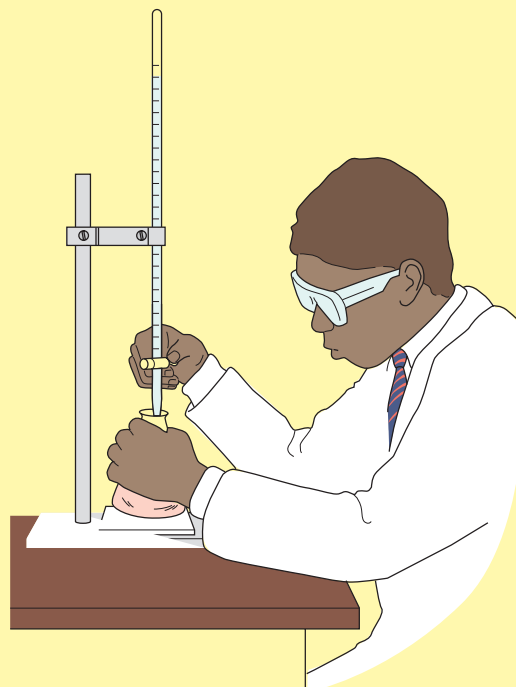**5** The acid is added until the alkali has been neutralised completely. This is shown by the blue colour of the indicator just disappearing.
6 The final reading on the burette at the end-point is recorded and further titrations carried out until consistent results are obtained (within 0.1 cm^3 of each other). Some sample data are shown on the right.

▲ **Figure 8.19** The titration is carried out accurately

Volume of sodium hydroxide solution = 25.0 cm^3

Average volume of 0.10 mol/dm^3 solution of hydrochloric acid added = 21.0 cm^3

The neutralisation reaction which has taken place is:

hydrochloric + sodium → sodium + water
    acid   hydroxide   chloride

$$HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$$
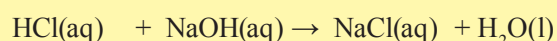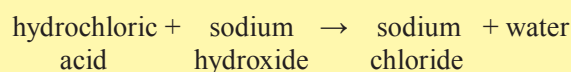
From this equation, it can be seen that 1 mole of hydrochloric acid neutralises 1 mole of sodium hydroxide.

Now you can work out the number of moles of the acid using the formula given in Chapter 4, p. 59.

$$moles = \frac{volume}{1000} \times concentration$$

$$= 21.0 \times \frac{0.10}{10000}$$

$$= 21.0 \times 10^{-3}$$

number of moles of hydrochloric acid = number of moles of sodium hydroxide

Therefore, the number of moles of sodium hydroxide = 2.1×10^{-3}
2.1×10^{-3} moles of sodium hydroxide is present in 25.0 cm^3 of solution.
Therefore, in 1 cm^3 of sodium hydroxide solution we have:

$$\frac{2.1 \times 10^{-3}}{25.0} \text{ moles}$$

Therefore, in 1 dm^3 of sodium hydroxide solution we have:

$$\frac{2.1 \times 10^{-3}}{25.0} \times 1000 = 0.084 \text{ mole}$$

The concentration of sodium hydroxide solution is 0.084 mol/dm^3.
You can simplify the calculation by substituting in the following mathematical equation:

$$\frac{M_1 V_1}{M_{acid}} = \frac{M_2 V_2}{M_{alkali}}$$

where:

M_1 = concentration of the acid used

V_1 = volume of acid used (cm^3)

M_{acid} = number of moles of acid shown in the chemical equation

M_2 = concentration of the alkali used

V_2 = volume of the alkali used (cm^3)

M_{alkali} = number of moles of alkali shown in the chemical equation.

In the example:

M_1 = 0.10 mol/dm^3

V_1 = 21.0 cm^3

M_{acid} = 1 mole

M_2 = unknown

V_2 = 25.0 cm^3

M_{alkali} = 1 mole

Substituting in the equation:

$$\frac{0.10 \times 21.0}{1} = \frac{M_2 \times 25.0}{1}$$

Rearranging:

$$M_2 = \frac{0.10 \times 21.0 \times 1}{1 \times 25.0}$$

$$M_2 = 0.084$$

The concentration of the sodium hydroxide solution is 0.084 mol/dm^3.

Another example of a titration calculation could involve a neutralisation reaction in which the ratio of the number of moles of acid to alkali is not 1 : 1. The example below shows how such a calculation could be carried out.

? Worked example

In a titration to find the concentration of a solution of sulfuric acid, 25 cm^3 of it was just neutralised by 20.15 cm^3 of a 0.2 mol/dm^3 solution of sodium hydroxide. What is the concentration of the sulfuric acid used?

First, write out the balanced chemical equation for the reaction taking place.

sulfuric acid + sodium hydroxide → sodium sulfate + water

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

From this balanced equation, it can be seen that 1 mole of sulfuric acid reacts with 2 moles of sodium hydroxide.

Therefore, the number of moles of sodium hydroxide used:

$$= 20.15 \times \frac{0.2}{1000} = 4.03 \times 10^{-3}$$

The number of moles of sulfuric acid which will react with 4.03×10^{-3} moles of sodium hydroxide:

$$= 4.03 \times 10^{-3} \times \frac{1}{2} = 2.015 \times 10^{-3}$$

This is the number of moles of sulfuric acid present in 25 cm^3 of the solution, so the concentration of the sulfuric acid is:

$$2.015 \times 10^{-3} \times \frac{1000}{25} = 0.081 \text{ mol/dm}^3$$

Test yourself

9 24.2 cm^3 of a solution containing 0.20 mol/dm^3 of hydrochloric acid just neutralised 25.0 cm^3 of a potassium hydroxide solution. What is the concentration of this potassium hydroxide solution?

10 22.4 cm^3 of a solution containing 0.10 mol/dm^3 of sulfuric acid just neutralised 25.0 cm^3 of a sodium hydroxide solution. What is the concentration of this sodium hydroxide solution?

Revision checklist

After studying Chapter 8 you should be able to:
✓ Describe the chemical properties of acids and bases.
✓ State that an alkali is a soluble base.
✓ Explain the difference between a concentrated and dilute acid.
✓ Describe the use of indicators to determine the presence of an acid or alkali and to determine their strengths.

✓ Explain the difference between a strong and weak acid.

✓ Explain what a neutralisation reaction is.
✓ Classify oxides as being acidic, basic or amphoteric .
✓ Describe the four different methods of making soluble salts.
✓ State the rules of solubility.

✓ Explain how an insoluble salt is made.
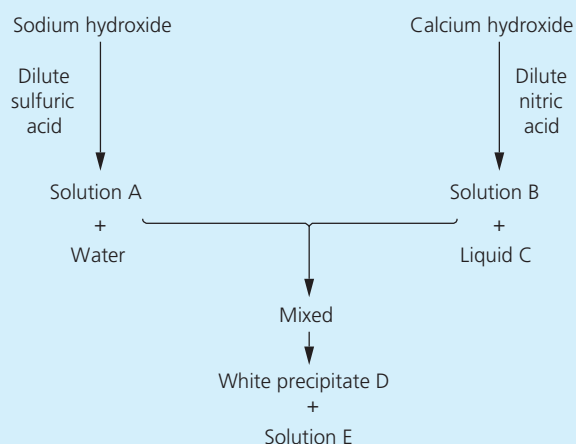
✓ Define the terms 'hydrated' and 'anhydrous'.

Exam-style questions

1 Explain, with the aid of examples, what you understand by the following terms:
   a strong acid [2]
   b weak acid [2]
   c strong alkali [2]
   d weak base. [2]

2 **a** Copy out and complete the table, which covers the different methods of preparing salts.

Method of preparation	Name of salt prepared	Two substances used in the preparation
Acid + alkali	Potassium sulfate	. . . . and . . . . [2]
Acid + metal	. . . . [1]	. . . . and dilute hydrochloric acid [1]
Acid + insoluble base	Magnesium sulfate	. . . . and . . . . [2]
Acid + carbonate	Copper . . . . . [1]	. . . . and . . . . [2]
Precipitation	Lead iodide	. . . . and . . . . [2]

   b Give word and balanced chemical equations for each reaction shown in your table. Also write ionic equations where appropriate. [15]

3 Consider the following scheme.

Sodium hydroxide   Calcium hydroxide

Dilute sulfuric acid   Dilute nitric acid

Solution A   Solution B
+   +
Water   Liquid C

Mixed

White precipitate D
+
Solution E

   a Give the names and formulae of substances **A** to **E**. [5]
   b Describe a test which could be used to identify the presence of water. [1]

   c Give the indicator suitable for the initial reaction between the hydroxides and the dilute acids shown. [1]
   d Give balanced chemical equations for the reactions taking place in the scheme. [3]
   e Give an ionic equation for the production of the white precipitate **D**. [1]

4 In a titration involving 24.0 cm³ potassium hydroxide solution against a solution of concentration 1 mol/dm³ of sulfuric acid, 28.0 cm³ of the acid was found to just neutralise the alkali completely.
   a Give a word and balanced chemical equation for the reaction. [2]
   b Identify a suitable indicator for the titration and state the colour change you would observe. [1]
   c Calculate the concentration of the alkali in mol/dm³. [3]
   d Describe a chemical test which you could use to identify the type of salt produced during the reaction. [2]

5 For each of the following, explain the term given **and** give an example:
   a neutralisation [2]
   b titration [3]
   c soluble salt [2]
   d insoluble salt. [2]

6 Copper(II) sulfate crystals exist as the *pentahydrate*, $CuSO_4.5H_2O$. It is a salt *hydrate*. If it is heated quite strongly, the *water of crystallisation* is driven off and the *anhydrous* salt remains.
   a Explain the meaning of the terms shown in italics. [4]
   b Describe the experiment you would carry out to collect a sample of the water given off when the salt hydrate was heated strongly. Your description should include a diagram of the apparatus used and a chemical equation to represent the process taking place. [4]
   c Describe a chemical test you could carry out to show that the colourless liquid given off was water. [2]

d Describe one other test you could carry out to show that the colourless liquid obtained in this experiment was pure water.　[1]

e Sometimes it is necessary to work out the percentage by mass of water of crystallisation as well as the number of moles of water present in a hydrated crystal.

　i Use the information given to calculate the percentage, by mass, of water of crystallisation in a sample of hydrated magnesium sulfate.　[3]

　　Mass of crucible = 14.20 g

　　Mass of crucible + hydrated $MgSO_4$ = 16.66 g

　　Mass after heating = 15.40 g

　ii Calculate the number of moles of water of crystallisation driven off during the experiment as well as the number of moles of anhydrous salt remaining.　[2]
　　(A_r: H = 1; O = 16; Mg = 24; S = 32)

　iii Using the information you have obtained in **ii**, write down, in the form $MgSO_4.xH_2O$, the formula of hydrated magnesium sulfate.　[1]