7.   Acids, Bases and Salts

7.1   Introduction

Many chemical compounds show properties which are characteristic of either acids, bases or salts. The concept of acids, bases and salts is therefore a very useful means for the classification of a large number of chemical compounds. In particular, the reactions between acids and bases to form salt and water, called **neutralization**, may be used in chemical analysis to determine the concentrations of some chemical substances in solution.

Furthermore, the control of the degree of acidity or alkalinity in most biological and chemical systems is very important for the proper functioning of such systems. For example, digestive enzymes in the stomach work best in dilute hydrochloric acid media whereas saliva as a digestive juice works best only in alkaline (basic) media.

Thus, a thorough understanding of the nature, general characteristics, reactions and methods of preparation of acids, bases and salts is essential in the study of chemistry.
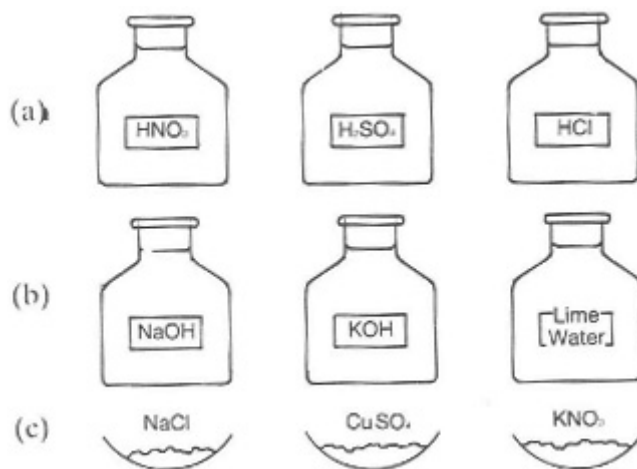

Figure 7.1 (a) acids; (b) bases; (c) salts.

ACIDS

7.2   Properties of acids

Many common substances are acids by nature. The sour tastes of unripe citrus fruits (orange, lime, grapefruit), are caused by their acid

contents. The sour taste of acids could be readily used to identify them. However, it is dangerous to taste chemical substances. The use of taste to identify acids is therefore not encouraged except in known food substances.

Acids show characteristic colour changes with certain dyes known as **indicators.** Thus, acids change the colour of wet litmus paper or litmus solution from blue to red. This is the commonest property of acids.

Experiment 7.1. Identifying Acidic Substances
You are given six different solutions containing:
(a) dilute solution of hydrogen chloride acid (hydrochloric acid)
(b) vinegar
(c) unripe orange juice
(d) dilute solution of sodium hydroxide
(e) solution of sodium chloride.

Test each solution with blue litmus paper by dipping half of the paper into the solution. You will observe that samples **a**, **b**, and **c** turned the blue litmus paper red, showing that they are acids. Samples **d** and **e** did not do this because they are not acids.

We have just learnt one way of distinguishing acids from other types of substances, using a simple property â€" that acids change the colour of litmus paper or solution from blue to red. Any substance like litmus which changes colour in the presence of an acid or base is called an **acid-base indicator.** Other common indicators which may be used to show the presence of an acid are methyl orange, phenolphthalein; and the universal indicator.

Experiment 7.2 Investigating the Action of Dilute Acid with Indicator
Into four test tubes labelled 1-4, place about 5cm of dilute (2 mol dm^{-3}) hydrogen chloride acid (hydrochloric acid), tetraoxosulphate(VI) acid, ethanoic acid (or vinegar), and lime juice respectively. Use little portions of each solution to carry out the following tests.
Test a portion of each solution with:
(i) litmus paper.
(ii) two drops of methyl orange.
(iii) two drops of phenolphthalein indicator.
(iv) two drops of universal indicator.
Record in tabular form as shown in Table 7.1 the colour of each indicator after it has been added to each solution.

TABLE 7:1

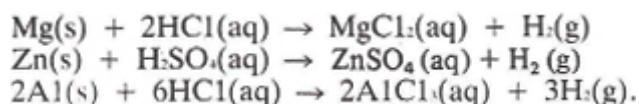|	Colour change of			
Acid Medium	Litmus	Methyl Orange	Phenolphthalein	Universal Indicator
Hydrogen chloride acid (Hydrochloric acid) Tetraoxosulphate (VI) acid Ethanoic acid Lime juice				

The action of acids on indicators is one of their characteristic properties. We shall proceed to investigate other characteristic properties of acids.

Experiment 7.3 Investigating the Action of Dilute Acids on Metals

Pour a little quantity of bench solutions of hydrogen chloride acid (hydrochloric acid), tetraoxosulphate(VI) acid, ethanoic acid and fresh unripe lime juice into four test-tubes labelled 1-4 and carry out the following tests.

(a) To the tube containing HCl, add a piece of magnesium ribbon of about 2 cm long. Rapid effervescence will be observed. Just before the magnesium has disappeared completely, apply a lighted splint over the mouth of the test-tube! Record your obvservation.

(b) Repeat (a) using each of the other acids in place of HCl.

(c) Repeat the procedures with fresh solutions of the acids, and pieces of the metals zinc, aluminium, copper and lead.

Effervescence will be observed only with magnesium, zinc and aluminium. It is rapid with some of the acids but slow in others. The equation for the reactions are:

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$
$$Zn(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2(g)$$
$$2Al(s) + 6HCl(aq) \rightarrow 2AlCl_3(aq) + 3H_2(g).$$

A pop sound is produced when a lighted splint is applied to the mouth of the test tube, showing that the gas liberated is hydrogen.

Experiment 7.4 Action of Dilute Acids on Trioxocarbonate(IV) and Hydrogen Trioxocarbonate(IV) Salts

Place some calcium trioxocarbonate(IV) into a test-tube, then add dilute hydrogen chloride acid (hydrochloric acid). Pass the gas evolved into a solution of calcium hydroxide (**lime water**) as shown in Figure 7.2.
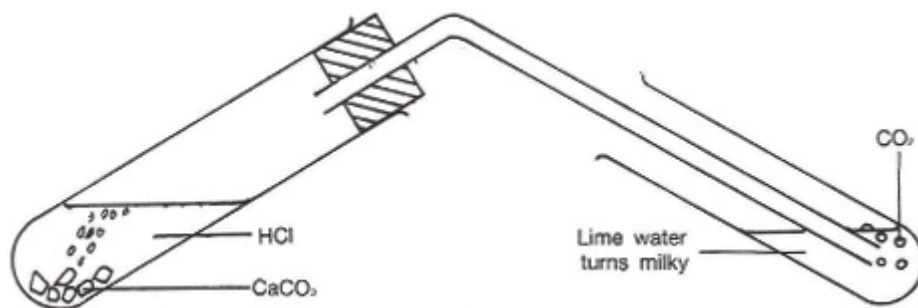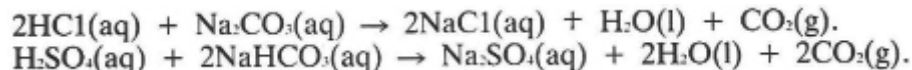
Figure 7.2 **Action of Carbon(IV) oxide on lime water**

Record your observations and write equations for the reaction.

Repeat the experiment using sodium hydrogen trioxocarbonate(IV), calcium hydrogen trioxocarbonate(IV) and magnesium trioxocarbonate(IV) in place of calcium trioxocarbonate(IV). Record your observations and write equations.

Add a little quantity of the following solutions; lime juice, ethanoic acid, unripe orange juice and dilute trioxonitrate(V) acid, to separate samples of sodium trioxocarbonate(IV) in test-tubes. Record your observations.

In each of these tests, effervescence will occur in the cold on the addition of **the dilute acid. The gas evolved turns lime water milky showing that it is** carbon(IV) oxide (CO_2). The equations for the reactions are:

$$2HCl(aq) + Na_2CO_3(aq) \rightarrow 2NaCl(aq) + H_2O(l) + CO_2(g).$$
$$H_2SO_4(aq) + 2NaHCO_3(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l) + 2CO_2(g).$$

Characteristics of acids

The characteristic reactions of acids are those shown by their dilute solutions.

1. They have sour tastes.
2. Strong acids are corrosive. The more concentrated the acid, the more corrosive it is.
3. Acids react with indicators to give characteristic colours. (Blue litmus â€" red; phenolpthalein â€" colourless; methyl orange â€" pink).
4. They react with metals above hydrogen in the activity series to liberate hydrogen. Dilute trioxonitrate(V) acid does not liberate hydrogen with metals.
5. They react with trioxocarbonate(IV) and hydrogen trioxocarbonate(IV) salts to liberate carbon(IV) oxide.
6. They are neutralized by bases.

7.3   Acids and hydroxonium ions

The reactions described above are exhibited by all acids. They have

been described as the characteristic or distinguishing properties of acids. There must be a fundamental factor common to all acids which explains these characteristic properties.

Generally, pure acids do not show the characteristic acidic properties and do not conduct electricity, but when dissolved in water they behave like acids and become conductors of electricity. For example, pure tetraoxosulphate(VI) acid does not liberate carbon(IV) oxide with dry solid calcium trioxocarbonate(IV) or liberate hydrogen with magnesium or zinc. Only the dilute solution of the acid undergoes these reactions: water is therefore essential for the manifestation of acidic behaviour. Pure acids are covalent compounds.

The formula of some common acids are shown in Table 7.2. Each contains at least one atom of hydrogen permolecule. This functional hydrogen atom is responsible for the characteristic reactions. We may therefore represent an acid by the general formula HA where H is the functional hydrogen atom and A an atom or a group of atoms of other element(s) (such as Cl^-, $HCOO^-$, NO_3, HSO_4^-) which are called acid radicals.

When the acid HA is dissolved in water, the water causes the acid to undergo *ionization or dissociation* which leads to the production of hydroxonium ions.

TABLE 7.2 Common aids and their molecular formulae

Name of Acid	Formula
Hydrogen chloride acid (Hydrochloric acid)	HCl
Tetraoxosulphate(VI) acid	H_2SO_4
Trioxonitrate(V) acid	HNO_3
Methanoic acid	H_2CO_2 (HCOOH)
Ethanoic acid	$H_4C_2O_2$ (CH$_3$COOH)
Elhanedioic acid	$H_2C_2O_4$

$$HA + H_2O(l) \longrightarrow H_3O^+ (aq) + A^- (aq)$$
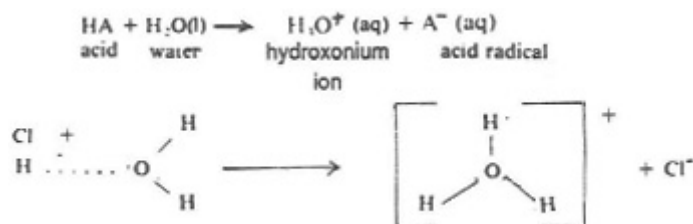
The hydroxonium ion results from the dative bond formed between the hydrogen ion (a proton) and a water molecule.

The ionisation reaction in water is a general reaction of all acids. Thus when any acid is dissolved in water hydroxonium ions are produced.

An acid is therefore defined as **any substance which when**

**dissolved in water produces hydroxonium ions.**

The ionisation reactions of some common acids are shown in Table 7.3.

TABLE 7.3 Dissociation of some common acids

Name of Acid	Formula	Basicity	Dissociation steps			
			acid	water	hydroxonium ion	acid radical
Hydrogen chloride acid (hydrochloric acid)	HCl	1 (monobasic)	$HCl + H_2O \longrightarrow H_3O^+ + Cl^-$ (chloride ion)			
Trioxo- nitrate (V) acid	HNO_3	1 (monobasic)	$HNO_3 + H_2O \longrightarrow H_3O^+ + NO_3^-$ (trioxonitrate(V) ion)			
Ethanoic acid	CH_3COOH	1 (monobasic)	$CH_3COOH + H_2O \longrightarrow H_3O^+ + CH_3COO^-$ (ethanoate ion)			
Tetraoxo- sulphate(VI) acid	H_2SO_4	2 (dibasic)	(i) $H_2SO_4 + H_2O \longrightarrow H_3O^+ + HSO_4^-$ (hydrogen tetraoxosulphate(VI) ion) (ii) $HSO_4 + H_2O \longrightarrow H_3O^+ + SO_4^{2-}$ (tetraoxosulphate(VI) ion)			
Ethanedioic acid	COOH | COOH $(H_2C_2O_4)$	2 (dibasic)	(i) $H_2C_2O_4 + H_2O \longrightarrow H_3O^+ + HC_2O_4^-$ (hydrogen ethanedioate ion) (ii) $HC_2O_4 + H_2O \longrightarrow H_3O^+ + C_2O_4^{2-}$ (eth. (ethanedioate ion)			
Tetraoxo- phosphate(V) acid	H_3PO_4	3 (tribasic)	(i) $H_3PO_4 + H_2O \longrightarrow H_3O^+ + H_2PO_4^-$ (dihydrogen tetraoxo- phosphate(V) ion) (ii) $H_2PO_4 + H_2O \longrightarrow H_3O^+ + HPO_4^{2-}$ (hydrogen tetraoxo phosphate(V) ion) (iii) $HPO_4^{2} + H_2O \longrightarrow H_3O^+ + PO_4^{3-}$ (tetraoxophosphate (V) ion)			

7.4   Basicity of acids

Although many acids especially the organic acids contain more than one hydrogen atom in their molecules, not all these hydrogen atoms are replaceable by a metal or can ionize in water. For example, one molecule of ethanoic acid (CH_3COOH) reacts with magnesium to liberate only one atom of hydrogen.

$$2CH_3COOH(aq) + Mg(s) \rightarrow Mg(CH_3COO)_2(aq) + H_2(g)$$
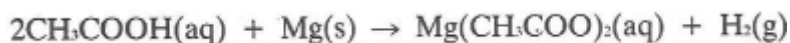
In some other acids however, two or three atoms of hydrogen per molecule of the acid are replaceable by a metal.

The number of replaceable hydrogen atoms in a molecule of an

acid is called the basicity of the acid.

Some common acids are grouped in Table 7.4 according to their basicities. For acids which are not monobasic, the hydrogen atoms are replaced one after the other, e.g.
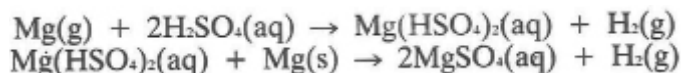
$$Mg(g) + 2H_2SO_4(aq) \rightarrow Mg(HSO_4)_2(aq) + H_2(g)$$
$$Mg(HSO_4)_2(aq) + Mg(s) \rightarrow 2MgSO_4(aq) + H_2(g)$$

TABLE 7.4 Basicity of some common acids

Monobasic (Basicity = 1)	Dibasic (Basicity = 2)	Tribasic (Basicity = 3)
Hydrogen chloride acid (Hydrochloric acid) HCl Trioxonitrate(V) acid, HNO_3 Methanoic acid, HCOOH Ethanoic acid, CH_3COOH	Tetraoxosulphate(VI) acid, H_2SO_4 Trioxosulphate(IV) acid, H_2SO_3 Trioxocarbonate(IV) acid, H_2CO_3 Ethanedioic acid, $H_2C_2O_4$	Trioxophosphate(III) acid, H_3PO_3 Tetraoxophosphate(V) acid, H_3PO_4

7.5 Strong and weak acids

Some acids when dissolved in water ionise completely into hydroxonium ions and their acid radicals while others undergo incomplete ionisation. Such **acids which undergo complete ionisation in water are called strong acids**. Examples are HCl, H_2SO_4 and HNO_3. **Acids which undergo partial ionisation in water are called weak acids**. These acids produce low concentrations of hydroxonium ions in solution. Examples of some weak acids include: HCOOH, CH_3COOH, H_2CO_3

7.6 Methods for the preparation of acids

Acids are prepared in the laboratory by the following general methods:

1. By dissolving an acidic oxide (acid anhydride) in water

Experiment 7.5 Preparation of trioxosulphate(IV) acid from sulphur(IV) oxide
Ignite a spatula-full of sulphur in a deflagrating spoon and lower it into a gas jar containing some water as shown in Figure 7.3. When the sulphur stops burning, remove the deflagrating spoon and cover the
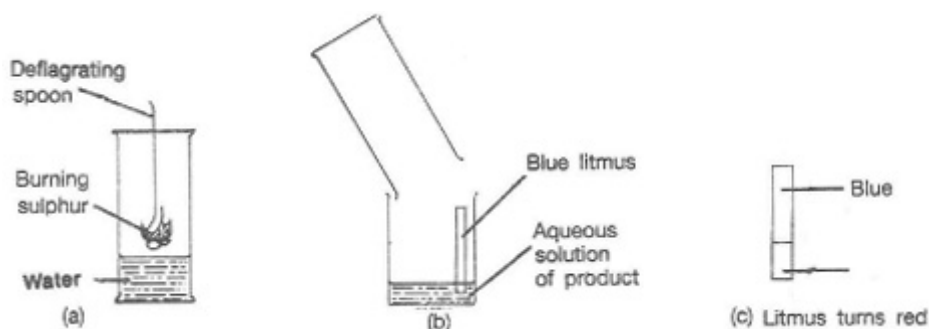
gas jar with its cover.

Figure 7.3 Preparation of an acid from its acid anhydride

Shake up the mixture in the gas jar. Transfer the solution into a beaker and dip a piece of blue litmus paper into it.

The litmus paper turns red showing that the solution is acidic. The sulphur(IV) oxide formed by the burning of sulphur in air reacts with water to form the acidic solution. The sulphur(IV) oxide is described as an acid anhydride.

The equations for the reactions that have taken place are,

$$S(s) + O_2(g) \rightarrow SO_2(g) \text{ (burning)}$$
$$SO_2(g) + H_2O(l) \rightarrow H_2SO_3(aq) \text{ (Trioxosulphate (IV) acid)}$$
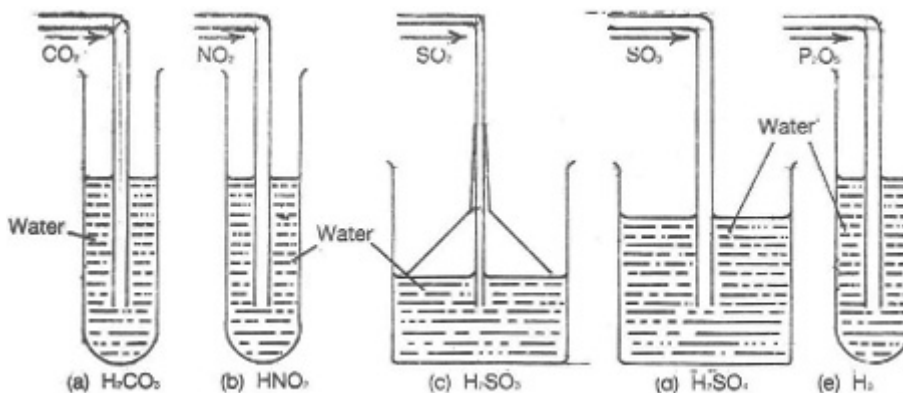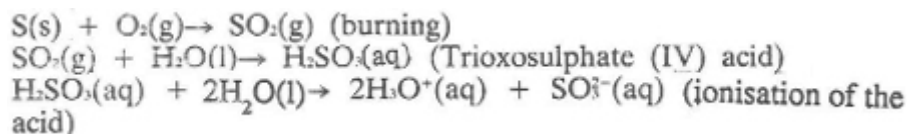$$H_2SO_3(aq) + 2H_2O(l) \rightarrow 2H_3O^+(aq) + SO_3^{2-}(aq) \text{ (ionisation of the acid)}$$

Figure 7.4 Preparation of acids from their acid anhydrides

Oxides of non-metals which dissolve in water to form acidic solutions (Figure 7.4) are called acid anhydrides.

The acid anhydrides are shown below.

TABLE 7.5

Oxide (Acid Anhydride)	Acid Formed
Carbon(IV) oxide	Trioxocarbonate(IV) acid
Nitrogen(IV) oxide	Dioxonitrate(III) acid

	and trioxonitrate(V) acid
Sulphur(IV) oxide	Trioxosulphate(IV) acid
Sulphur(VI) oxide	Tetraoxosulphate(VI) acid
Phosphorus(V) oxide	Tetraoxophosphate(V) acid

Exercise 7A

(a) Write equations for the formation of the above named acid anhydrides.

(b) Write balanced equations for the formation of the corresponding acids from the following acid anhydrides:
CO_2, NO_2, SO_2, SO_3, P_2O_5

2. By the displacement of weaker and/or more volatile acids from their salts

Experiment 7.6 Preparation of Hydrogen Chloride Acid (Hydrochloric Acid) from Sodium Chloride

Set up a round-bottomed flask as shown in Figure 7.5. A delivery tube and a thistle funnel are fitted to the cork of the flask. Lower the inverted funnel attached to the other end of the delivery tube slightly below the surface of water contained in a beaker.
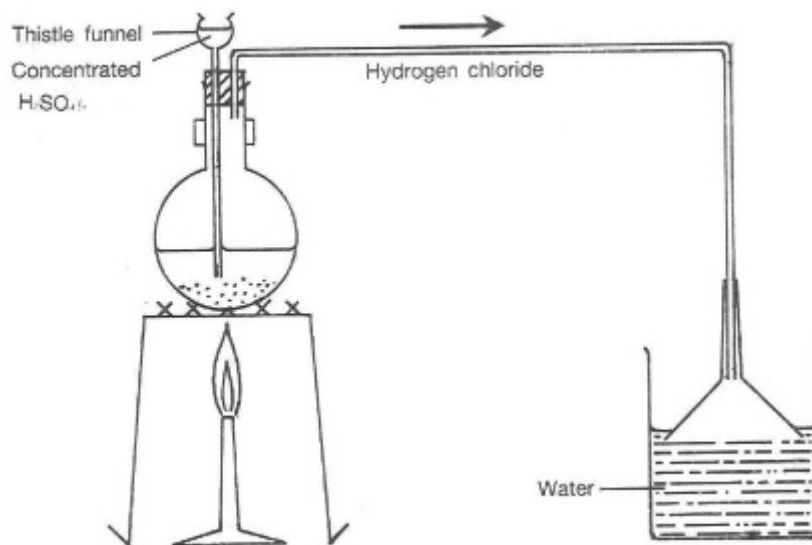

Figure 7.5 Preparation of hydrogen chloride acid (hydrochloric acid) by displacement from sodium chloride

Place some sodium chloride in the flask and then, slowly add some concentrated tetraoxosulphate(VI) acid from the funnel. Gently heat the mixture and pass the liberated hydrogen chloride gas into water through the inverted funnel.

(a) Test the solution with blue litmus paper.

In this preparation, the concentrated tetraoxosulphate(VI) acid

displaces the more volatile hydrogen chloride gas. The gaseous hydrogen chloride dissolves in water to give hydrogen chloride acid (hydrochloric acid).

$$NaCl(s) + H_2SO_4(aq) \xrightarrow{heat} NaHSO_4(aq) + HCl(g).$$
$$HCl(g) + H_2O(l) \rightarrow H_3O^+(aq) + Cl^-(aq).$$
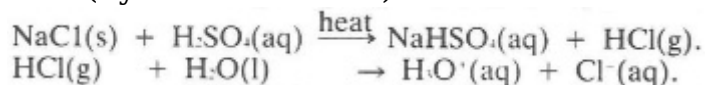
The hydrogen chloride is passed into the water in the beaker through the funnel because it is very soluble in water. If it is delivered through a delivery tube water will be sucked back into the flask because of the very high solubility of the gas.

Other acids which may be prepared by the displacement method using tetraoxosulphate(VI) acid include:

(i) Trioxonitrate(V) acid

(ii) Trioxocarbonate(IV) acid

(iii) Methanoic acid

(iv) Ethanoic acid

A common feature of the method of preparation of each of the above acids is that either:

(a)   a volatile acid is displaced by a less volatile acid, as in the preparations of hydrogen chloride acid (hydrochloric acid) and trioxonitrate(V) acid, or

(b)   a weak acid is displaced by a stronger acid, as in the preparations of trioxocarbonate(IV) acid, methanoic acid, and benzoic acid.

This method is used in the preparation of nearly all weak acids. (Note that benzoic acid is a solid organic acid.)

3. By double decomposition reactions

Experiment 7.7 Preparation of Trioxonitrate (V) acid

Your teacher will set up a Kipp's apparatus to generate hydrogen sulphide gas in a fume chamber.

Bubble the gas into three different test-tubes or beakers containing:

(i) a solution of lead(II) trioxonitrate(V)

(ii) a solution of lead(II) ethanoate, and

(iii) water.

Filter off the greyish-black precipitate formed in the first two test tubes and collect the filtrate in a beaker.

Test all three solutions with

(a) litmus paper, and

(b) methyl orange.

Comment on your observations.

In this preparation (Figure 7.6) the lead ions of the lead salt are removed as the insoluble salt (PbS) while the acid radical present remains in solution with hydroxonium ions derived from the hydrogen of the sulphide ($H_2 S$).
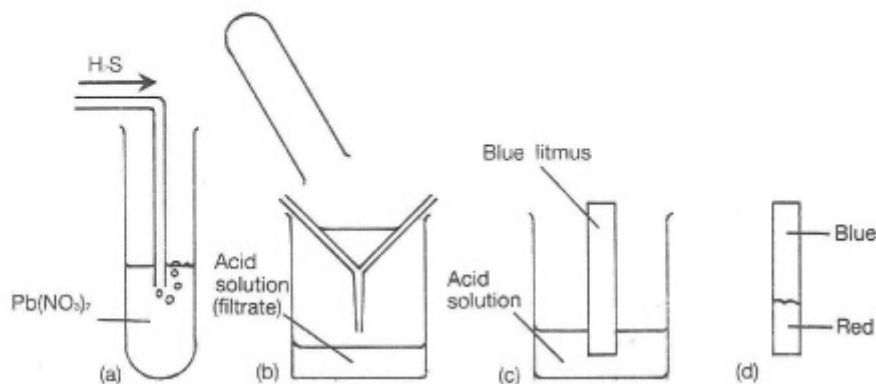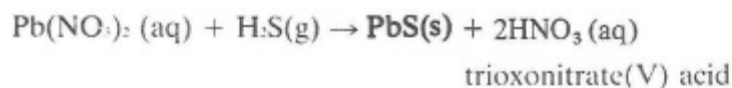
Thus:

$$Pb(NO_3)_2 \ (aq) + H_2S(g) \rightarrow \textbf{PbS(s)} + 2HNO_3 \ (aq)$$
$$\text{trioxonitrate(V) acid}$$

Figure 7.6 Preparation of acids by double decomposition

7.7 Uses of acids

The various uses to which acids could be put are illustrated in Figure 7.7
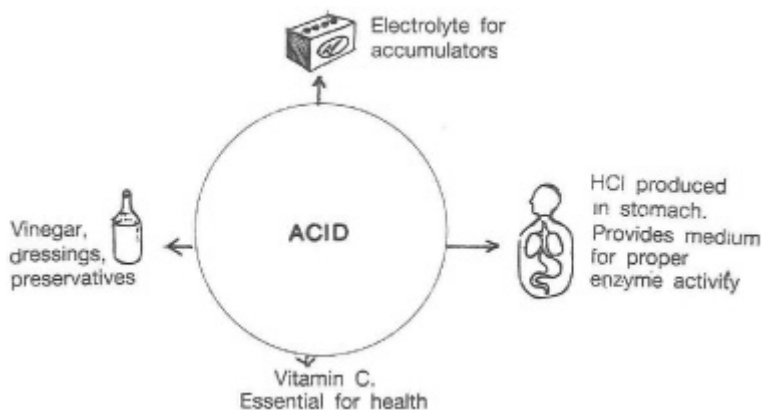

Figure 7.7 Uses of acids

BASES
7.8 Properties of bases

Bases form another important class of chemical substances. They are related to the acids in a somewhat opposite sense.

Tests for bases

Like acids, bases may be distinguished from other substances by their tastes because they have bitter tastes. Also, solutions of bases are slippery when rubbed between the fingers.
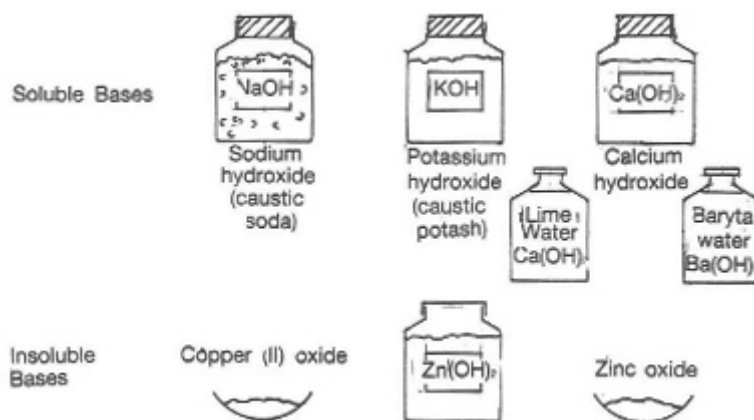

Figure 7.8 Varieties of bases

Experiment 7.8 Testing for bases with indicators

Prepare dilute solutions of the following substances, (i) Sodium hydroxide (ii) Potassium hydroxide (iii) Ammonia (iv) Sodium chloride, and (v) Magnesium oxide, and label the test tubes containing them. In another labelled test tube containing water, put in a spatula full of copper(II) oxide.

Test portions of the contents of each test-tube with

   (a)   red and blue litmus paper
   (b)   methyl orange (two drops)
   (c)   phenolphthalein (two drops)
   (d)   universal indicator.

Record your observations.

You would have observed that the solutions of sodium hydroxide, potassium hydroxide and ammonia turned red litmus blue.

In general, bases give characteristic colours with litmus, methyl orange, phenolphthalein and universal indicator as indicated below.

Indicator	Colour in base
Litmus	blue
Methyl orange	yellow
Phenolphthalein	yellow
Universal indicator	green to violet

Some bases, such as copper(II) oxide, which are insoluble in water may not show some of the basic properties discussed above. However, all bases undergo neutralization reactions.

Neutralization reaction

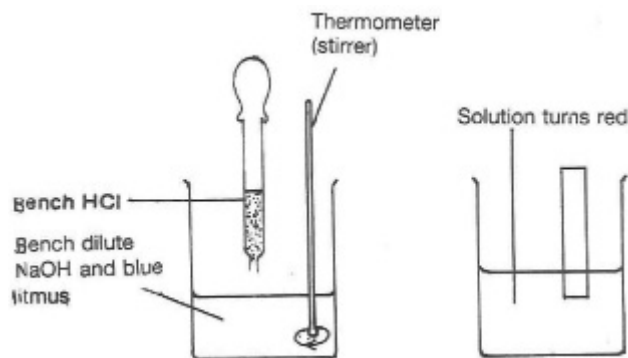*Experiment 7.9 Neutralization of an acid by a base*

Figure 7.9 Neutralization of an acid by a base

Pour about 5 cm³ of bench dilute sodium hydroxide solution into a small beaker. Add 2 drops of litmus solution. Record the colour of the solution after mixing. Place a thermometer in the mixture and record the temperature. Then **slowly add bench dilute hydrogen chloride acid (hydrochloric acid). You may use a burette or dropping** pipette to deliver the acid drop by drop. While adding the base, stir the solution gently with the thermometer. Note the colour and temperature of the mixture after each addition. Continue the addition of dilute hydrogen chloride acid (hydrochloric acid) and stirring until a colour change appears. Record the colour change and the temperature of the solution.
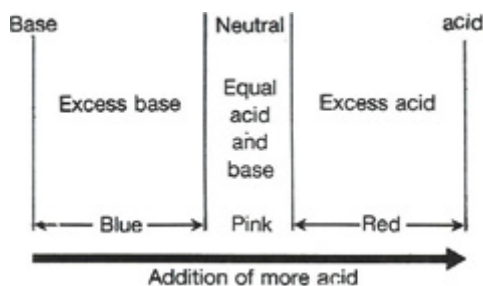

Figure 7.10 Colour changes of litmus indicator on addition of HCl and NaOH solutions

The solution is blue in colour when the litmus is first added because the solution is strongly basic. As the dilute hydrogen chloride acid (hydrochloric acid) solution is added, the colour of the solution remains blue, but the temperature rises, showing that a chemical change is taking place between the acid and the base. At some point during the addition, all the acid added is just enough to react with all the base in the solution. At this point the colour of the solution turns pink. When this happens, neutralization is said to have occured. It the point when neutralization occurs, the amounts of the acid and the base

present have satisfied the ratio in the balanced equation:

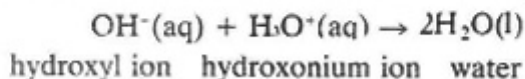$$HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$$
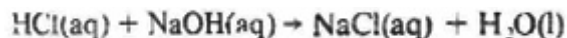
Neutralization is defined as a reaction between an acid and a base to form a salt and water only.

In the above example, the new solution contains sodium chloride and water only.

If we write the ionic form of the above neutralization equation, we notice that both the sodium (Na^+), and chloride (Cl^-) ions remain unchanged in solution before and after the reaction. They are not involved in the neutralization reaction, and are called **spectator** ions.

$$H Cl (aq) + Na OH (aq) \rightarrow Na^+(aq) + Cl^-(aq) + H_2O(l)$$
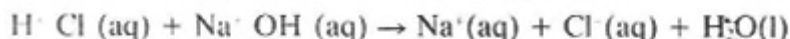
Therefore, the essential reaction in the neutralization process is:

$$OH^-(aq) + H_3O^+(aq) \rightarrow 2H_2O(l)$$

hydroxyl ion   hydroxonium ion   water

This ionic equation is used to represent all neutralization reactions involving a hydroxide and an acid.

All bases are capable of neutralizing solutions of acids.

Experiment 7.10 Neutralization reactions of bases
1. Place about 5cm³ of bench dilute hydrochloric acid in a test tube. Add two drops of litmus solution.

   Add from a watch glass samples of zinc oxide into the beaker (Figure 7.11). Stir the mixture with a glass rod after each addition. Continue the addition till there is a colour change.

   Repeat the procedure above using solid samples of sodium hydroxide, calcium oxide, magnesium oxide barium hydroxide, potassium hydroxide and zinc hydroxide.

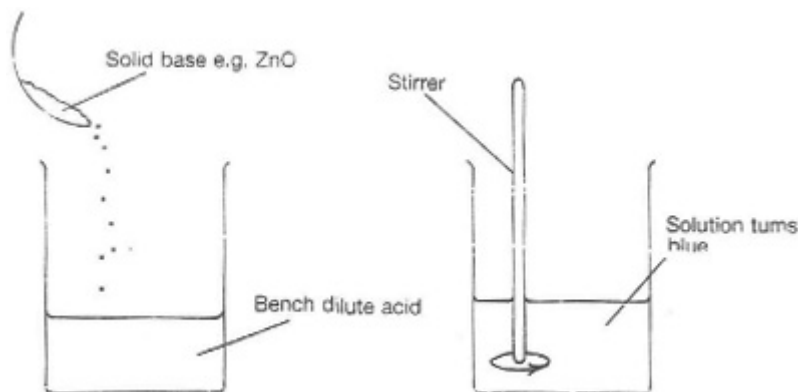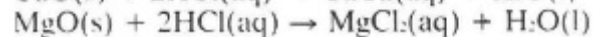   Does the sample eventually neutralize the acid solution? What is the chemical nature of the sample?
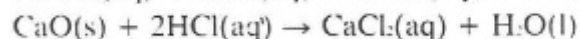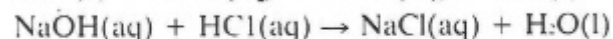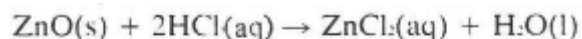

Figure 7.11 Neutralization reactions of bases

All the samples investigated in the above experiment turn the litmus solution pink. They neutralize the acid and therefore they are all bases.

The equations for the reactions are:

$$ZnO(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2O(l)$$
$$NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$$
$$CaO(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l)$$
$$MgO(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2O(l)$$

The common bases are either oxides or hydroxides of metals.

In general a base is defined as a hydroxide or metallic oxide which neutralizes an acid to form a salt and water.

7.9   Classification of bases

Experiment 7.11 Soluble and insoluble bases
Place some water in a test tune. Using the edge of a spatula, add a little sample of solid sodium hydroxide. Stir up with a glass rod. Is the sample soluble in water?

Repeat the solubility test on calcium oxide, copper(II) oxide, magnesium oxide, barium hydroxide, zinc hydroxide, potassium hydroxide and any other available metallic oxide or hydroxide. List those bases which are soluble and those which are insoluble.

Bases can be classified according to their chemical composition and **solubility** in water. Such a classification of the common bases is shown in Table 7.6.

TABLE 7.6 Soluble and insoluble bases

Chemical nature	Soluble Bases	Insoluble Bases
Oxides	Na_2O K_2O	MgO, PbO, FeO, Fe_2O_3 CuO, ZnO
Hydroxides	KOH $NaOH$ $Ba(OH)_2$ } slightly $Ca(OH)_2$ } soluble	$Zn(OH)_2, Pb(OH)_2$ $Cu(OH)_2,$ $Fe(OH)_2$ $Fe(OH)_3, Al(OH)_3$

As a general rule, all bases (oxide and hydroxide) of Group 1 metals in the Periodic table, with the exception of lithium, are soluble in water. Bases of calcium and barium are slightly soluble. All other bases are insoluble in water.
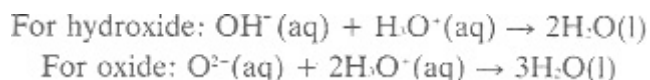
Bases which are soluble in water give basic aqueous solutions. They form a distinct class of bases called Alkalis. **An alkali is a soluble base;** the aqueous solution of a soluble base is said to be **Alkaline**.

Alkalis as a class consist of only a few common bases, namely: Potassium hydroxide, sodium hydroxide, barium hydroxide, calcium hydroxide, and ammonia solution.

Alkalis dissolve in water to produce hydroxide (hydroxyl) ions (OH⁻). For example:

$$Na_2O(s) + H_2O(l) \rightarrow 2NaOH(aq)$$

$$NaOH(s) \xrightarrow{\text{water}} Na^+(aq) + OH^-(aq)$$
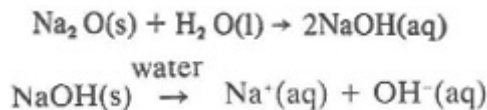
The hydroxide and oxide ions, OH⁻ and O²⁻ are responsible for the characteristic reactions of common bases. The reactions of bases with acids may be represented by the two equations:

$$\text{For hydroxide: } OH^-(aq) + H_3O^+(aq) \rightarrow 2H_2O(l)$$
$$\text{For oxide: } O^{2-}(aq) + 2H_3O^+(aq) \rightarrow 3H_2O(l)$$

Based on our studies so far, we can now summarise the characteristic properties of bases as follows:
1. Bases neutralize acids to form salt and water only.
2. Alkalis change the colour of litmus from red to blue.
3. Alkalis are slippery when rubbed between the fingers.
4. Alkalis have bitter taste.
5. Alkalis are caustic to the skin.
6. Alkalis (with the exception of ammonia solution) liberate ammonia when warmed with ammonium salts.

Caustic and Mild Alkali

The two best known alkalis — sodium hydroxide and potassium hydroxide have corrosive actions on the skin. Consequently they are often referred to as 'caustic alkalis', and hence called caustic soda and caustic potash respectively.

Some soluble trioxocarbonate(IV) salts, for example sodium trioxocarbonate(IV), on dissolution in water, show the characteristics of alkalis. Their solutions turn litmus from red to blue. Like bases, they neutralic acids, but in addition liberate carbon(IV) oxide, according to the equation.

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

The solutions of these trioxocarbonate(IV) salts though alkaline in behaviour, are not corrosive to the skin. They are therefore occasionally referred to as **mild alkalis**.

Caustic alkalis	Mild alkali

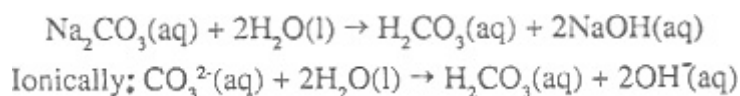| Caustic soda, NaOH | Sodium trioxocarbonate (IV), Na_2CO_3 |
| Caustic potash, KOH | Potassium trioxocarbonate (IV), K_2CO_3 |

It is important for us to recognise that 'mild alkalis' are actually salts and not true bases or alkalis.

The alkaline behaviour of these salts is due to a phenomenon called 'hydrolysis'.

Hydrolysis is a reaction in which a portion of a salt dissolves in water to form an acid and a base.

One of these, acid/base, is strong while the other is weak. The weak one remains unionised while the strong one is ionised. The ionised one shows its characteristic properties.

Sodium trioxocarbonate(IV) for example, reacts with water to form trioxocarbonate(IV) acid which is essentially unionised in water. Some sodium hydroxide is also formed, which, being a strong base is completely ionised into Na^+ and OH^-. The OH^- gives the solution the characteristic alkaline reaction.

$$Na_2CO_3(aq) + 2H_2O(l) \rightarrow H_2CO_3(aq) + 2NaOH(aq)$$

$$\text{Ionically: } CO_3^{2-}(aq) + 2H_2O(l) \rightarrow H_2CO_3(aq) + 2OH^-(aq)$$

Iron(III) chloride, on the other hand, gives an acidic solution. It dissolves in water forming HCl a strong acid which ionises completely, and $Fe(OH)_3$, a weak base that is almost unionised. The H^+ from HCl forms H_3O^+ with water to give the solution an acidic character. In general, salts of strong acids and weak bases give an acidic solution while salts of weak acids and strong bases give alkaline solutions.

Some properties of common bases are given in Table 7.7

TABLE 7.7 Physical Properties of Common Bases

Name	Physical State	Solubility in Water
Potassium hydroxide Sodium hydroxide	white solid	very soluble and colourless solution
Barium hydroxide Calcium hydroxide	white solid	slightly soluble and colourless solution
Magnesium oxide		insoluble
Zinc oxide	white solid	insoluble
Zinc hydroxide		insoluble
Diiron (II) trioxide	black solid	insoluble
Iron (II) hydroxide	green solid	insoluble
Iron (III) hydroxide	brown solid	insoluble
Lead (II) oxide	yellow solid	insoluble
Copper (II) oxide	black solid	insoluble
Copper (II) hydroxide	blue solid	insoluble

7.10 General methods of preparing bases

Bases can be prepared by the following methods:

1. By direct reaction between metals and oxygen:
   The reaction may be carried out by either of these two methods:

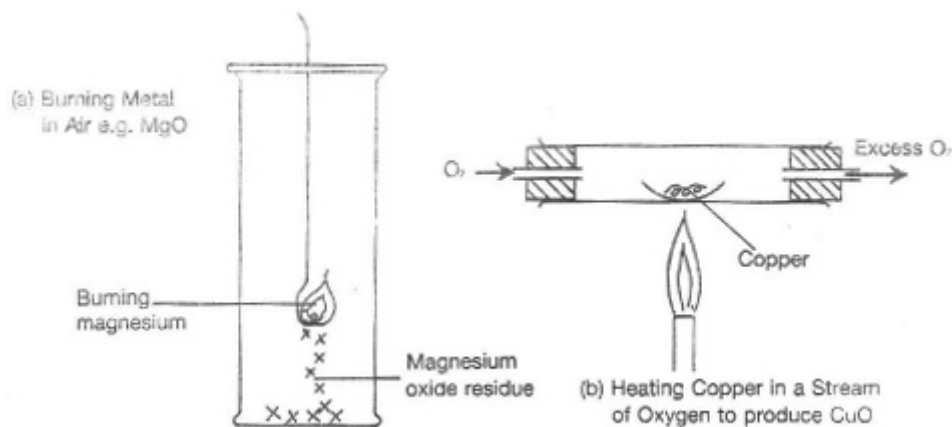(a) Burning the metal in air. Sodium, calcium and magnesium burn in air to form basic oxides.

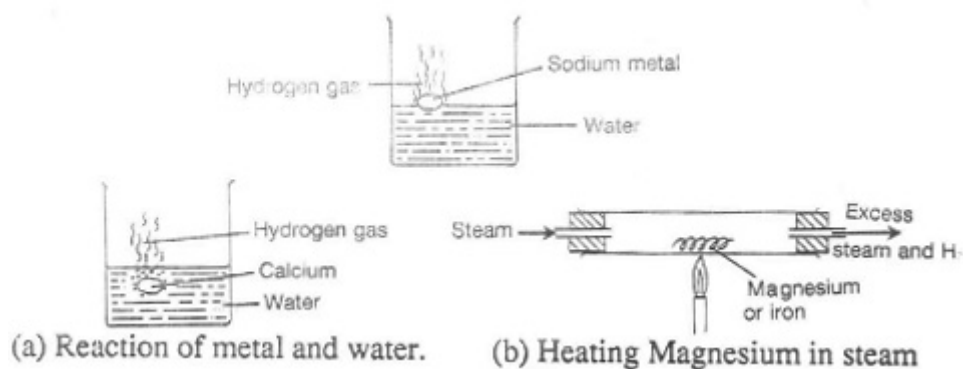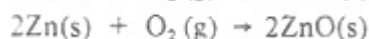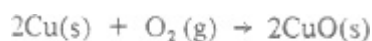**Figure 7.12 Preparation of bases by burning metals in oxygen**

(a) Reaction of metal and water.   (b) Heating Magnesium in steam

Figure 7.13 Preparation of bases by the action of metals on water (steam): NaOH, Ca(OH)$_2$, Mg(OH)$_2$

(b) Heating the metal in air or in a stream of oxygen.
Example:

$$2Cu(s) + O_2(g) \rightarrow 2CuO(s)$$
$$2Zn(s) + O_2(g) \rightarrow 2ZnO(s)$$

2. By the reaction between a metal and water:

(a) Reaction with cold water:
   The very reactive metals: potassium, sodium and calcium react with cold water to form alkalis (Figure 7.13).

Experiment 7.12 Preparation of sodium hydroxide from sodium metal
Half-fill a trough with distilled water and place it in a fume cupboard. Pick up a small piece of sodium with a pair of tongs and wipe it dry on a filter paper, then drop it into the water in the trough. Allow the sodium to dart about freely in the trough to avoid an explosion. When the reaction has stopped, test the resultingsolution with a piece of red litmus paper and record your observation. Dip a finger in the solution and rub it on another finger.
What do you feel?
The equation of the reaction is:

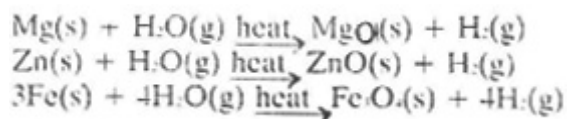$$2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$$
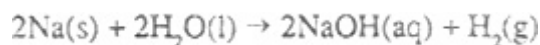
   You have just prepared a solution of sodium hydroxide. To obtain a concentrated solution from it, heat the resulting solution on a wire gauze until about one quarter of the original solution remains. This is a concentrated solution.
   The reaction between calcium metal and water to produce calcium hydroxide is milder than that of sodium. The product of this reaction is calcium hydroxide.

$$Ca(s) + 2H_2O(l) \rightarrow Ca(OH)_2(aq) + H_2(g)$$

Only soluble bases (alkalis) are prepared by this method,

(b) Reaction with steam: Metals which react with steam to yield oxides include magnesium, zinc and iron. The equations for their reactions are:

$$Mg(s) + H_2O(g) \xrightarrow{heat} MgO(s) + H_2(g)$$
$$Zn(s) + H_2O(g) \xrightarrow{heat} ZnO(s) + H_2(g)$$
$$3Fe(s) + 4H_2O(g) \xrightarrow{heat} Fe_3O_4(s) + 4H_2(g)$$

Note that copper and lead do not react with steam.

Experiment 7.13 Preparation of triirontetraoxide

Place some iron fillings in a porcelain boat. Carefully examine and record the physical appearance of the sample.

Introduce the boat into a combustion tube which is clamped horizontally and connect up the apparatus as shown in Figure 7.14.
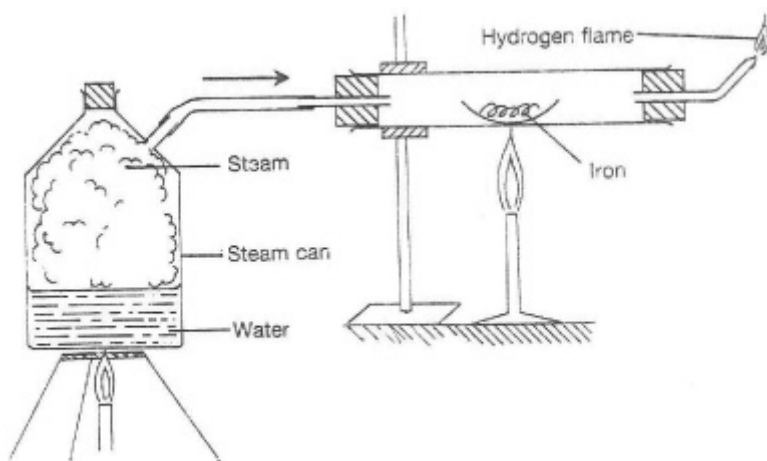

Figure 7.14 Preparation of triirontetraoxide

Generate steam from a steam can or a round-bottomed flask and pass it over the iron in the combustion tube which is heated very strongly. The hydrogen produced should be burnt at the jet as shown.

After about one hour, stop all heating and allow the apparatus to cool. Then remove the boat and examine its contents again. The black product is triiron tetraoxide, an insoluble base.

3. By thermal decomposition of oxygen – containing compounds of metals:

The trioxonitrate(V), trioxocarbonate(IV) and hydroxides of some metals are decomposed by heat to give the corresponding metal oxides. This method can be used to prepare the oxides of calcium, magnesium, iron, copper, zinc, aluminium and lead.
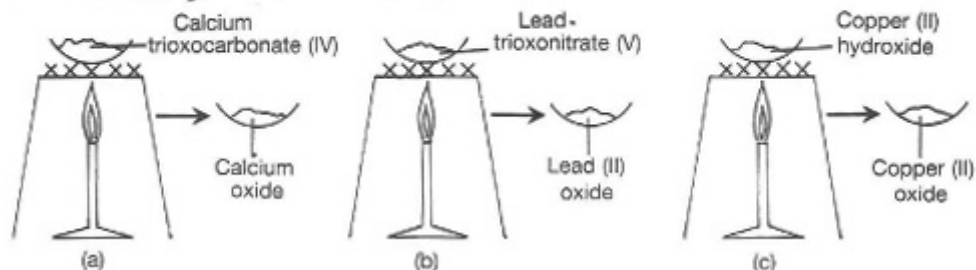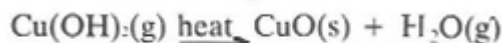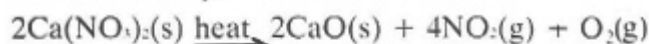
$$CaCO_3(s) \xrightarrow{heat} CaO(s) + CO_2(g)$$

$$2Ca(NO_3)_2(s) \xrightarrow{heat} 2CaO(s) + 4NO_2(g) + O_2(g)$$

$$Cu(OH)_2(g) \xrightarrow{heat} CuO(s) + H_2O(g)$$

Figure 7.15 Preparation of bases by thermal decomposition

Note that this method cannot be used to prepare potassium and sodium oxides becuse their trioxocarbonates(IV) and hydroxides are not decomposed by heat, while their trioxonitrates(V) give dioxonitrates(III) on heating.

4. By precipitation with a soluble hydroxide:

The method of precipitation may be used for the preparation of insoluble or sparingly soluble hydroxides such as those of calcium, zinc, magnesium, aluminium, iron, copper and lead.

Experiment 7.14 Preparation of hydroxides by precipitation

Place solutions of the trioxonitrates(V) of calcium, zinc, magnesium, aluminium, iron(II), iron(III), copper(II) and lead(II) into labelled test-tubes. as shown in Figure 7.16
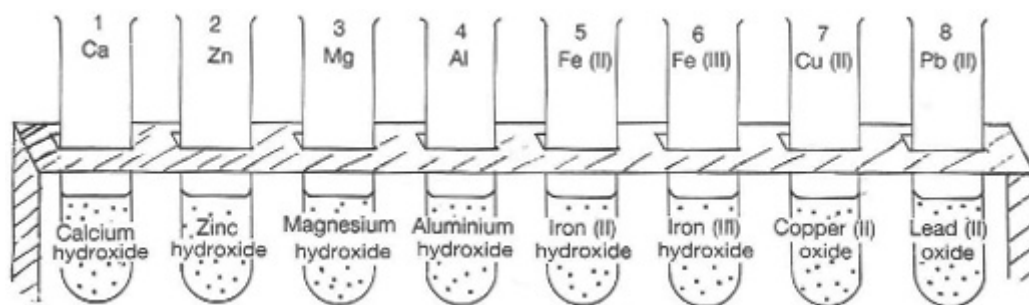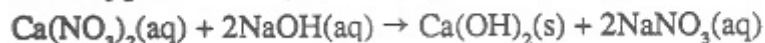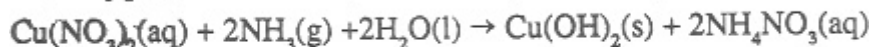

Figure 7.16 Preparation of insoluble hydroxides by precipitation

To each test-tube add bench dilute sodium hydroxide or ammonia solution in drops. Record the colours of the precipitates formed. Continue the addition of the alkali till no further preciptates are formed. (Note that ammonia solution does not precipitate calcium hydroxide. Can you explain this?)

Filter the mixtures and wash the precipitates with distilled water. Allow the residues to dry in-between two filter papers. Finally dry in an oven at low temperature.

Equations for the precipitation reactions are:

$Fe(NO_3)_2(aq) + 2NaOH(aq) \rightarrow Fe(OH)_2 + 2NaNO_3(aq)$

$Cu(NO_3)_2(aq) + 2NH_3(g) + 2H_2O(l) \rightarrow Cu(OH)_2(s) + 2NH_4NO_3(aq)$

$Ca(NO_3)_2(aq) + 2NaOH(aq) \rightarrow Ca(OH)_2(s) + 2NaNO_3(aq)$

Ionically

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

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

$Fe^{2-}(aq) + \quad 3OH^-(aq) \rightarrow Fe(OH)_3(s)$
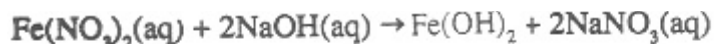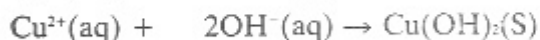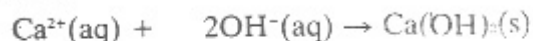
The colours of the precipitates of metal hydroxides are very important for the identification of the metals in qualitative analysis of the common metals. Hydroxides of iron and copper are coloured precipitates while the others are white.

5. By displacement of a volatile base by less volatile alkalis.

This is a special method for the preparation of ammonia, a volatile base.

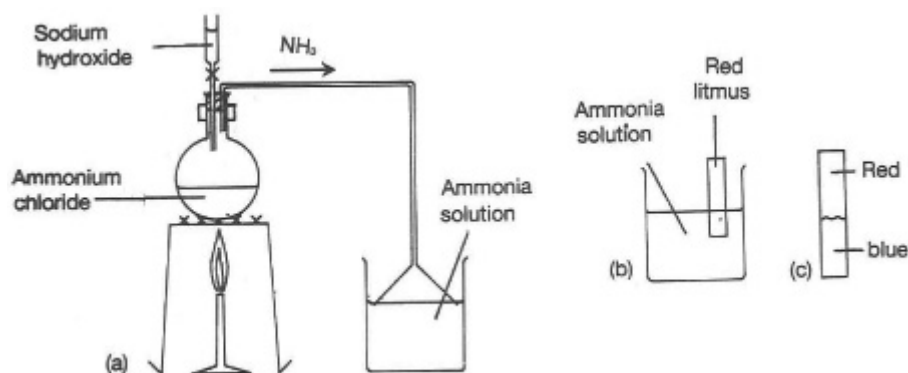*Experiment 7.15 Preparation of ammonia solution*

Figure 7.17 Displacement of a volatile base (ammonia) by strong alkali

Set up the apparatus shown in Figure 7.17. Place about four spatula loads of ammonium chloride into the round-bottomed flask. Pour some bench dilute sodium hydroxide through the tap funnel into the flask. Warm the flask over a bunsen flame. Ammonia gas is liberated and it passes through the inverted funnel into water in a beaker, dissolving to form ammonia solution.

$NH_4Cl(aq) + NaOH(aq) \rightarrow NH_3(g) + NaCl(aq) + H_2O(l)$
$NH_3(g) + H_2O(l) \rightarrow NH_4^+(aq) + OH^-(aq)$
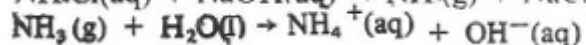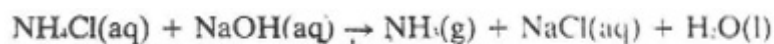
Any ammonium salt and any alkali, including calcium hydroxide, can be used for the preparation of ammonia gas.

Test the solution in the beaker with red litmus to confirm that it is basic.

Uses of Bases

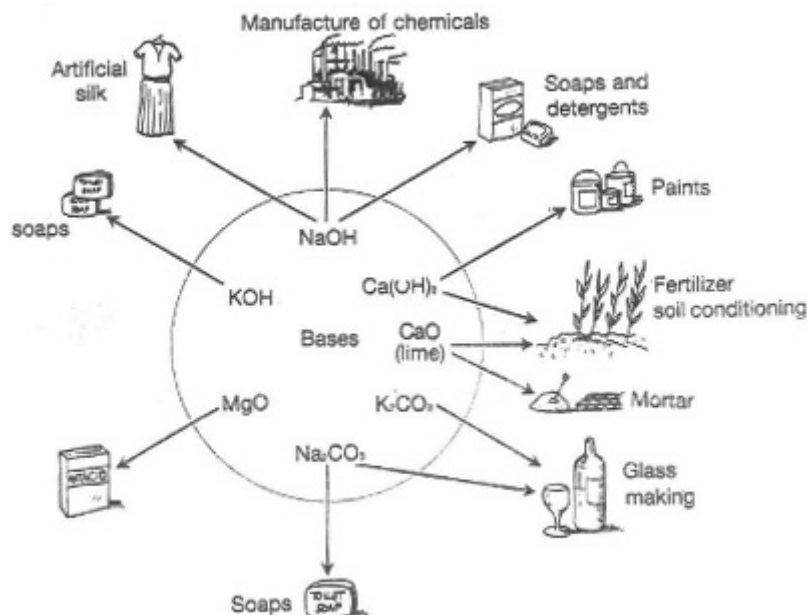The commonest uses of bases are illustrated in Figure 7.18.

Figure 7.18 Uses of bases

7.11   Acid-base indicators

In several of the previous experiments, litmus solution or paper was used to indicate whether a solution is acidic or basic. In the discussion of neutralization we observed that in the presence of excess acid, the litmus solution or paper shows a red colour. When the solution is neutral the colour is pink, but in the presence of excess base (alkali), the litmus solution is blue.

We learnt earlier that a substance which shows different colours in acidic and basic solutions is called an indicator. Most acid-base indicators are organiccompounds which behave as weak acids. Their chemical nature and colour in a given solution depends on whether or not they lose a proton or remain as the unionised molecule.

The general reaction of such an acid — base indicator is represented by the equation;

$$\underset{\text{(acid colour)}}{HIn\ (aq) + H_2O(l)} \rightleftharpoons \underset{\text{(base colour)}}{H_3O^+(aq) + In^-(aq)}$$

The anion (represented by In^-) has a different colour from the acid molecule, HIn.

In the presence of alkali which produce hydroxide ions, the indicator which is a weak acid, HIn, ionizes to form a large concentration of In^- which gives the base colour of the indicator. But in the presence of acidic solutions the hydroxonium ions from the acid prevent the ionization of the indicator so that it remains in the molecular form,

HIn, and thus manifests the acid colour.

The most commonly used indicators are litmus, methyl orange, phenolphthalein and methyl red. Their colour changes are shown below.

Indicator	Colour in acid	Colour in alkali
Litmus	Red	Blue
Methyl orange	Red	Yellow
Phenolphthalein	Colourless	Pink
Methyl red	Red	Yellow

Litmus indicator is a dye obtained from parts of plants by extraction with methanol or ethanol.

The indicator paper is made by dipping a filter paper into the dye solution and allowing it to dry. The most commonly used is the universed indicator paper.

A variety of other acid â€" base indicators could be readily obtained from parts of plants as in Experiment 7.16.

Experiment 7.16 Preparation of Acid â€" Base Indicators from Plants

Collect flowers of the following plants: *Hibiscus, Ixora, Pride of Barbados, Frangipani* (flowers of most other plant species may also be used).

Crush each flower in a mortar. Place about 20 â€" 30 cm³ of ethanol (or methanol) into the mortar; stir the crushed material into the ethanol, then decant and filter the solution as shown in Figure 7.19.

Collect the filtrate which is intensely coloured with the plant dye.
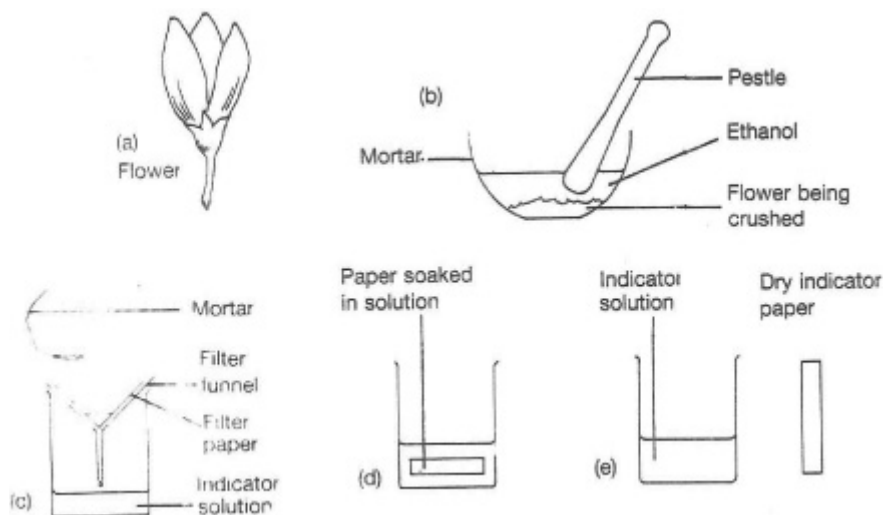

Figure 7.19 Preparation of indicator paper and solution

Prepare a table of results as shown below in Table 7.8.

Test the colours of the dyes in the named solution by placing two

drops of your indicator into each tube containing solution. Shake it to mix the contents. Record your observations and comment on your results.

TABLE 7.8 Colours of flower extracts in acidic and basic solutions

Extract from flower of	Colour in				
	Dilute HCl	Dilute NaOH	Dilute NaCl	Dilute Ethanoic acid	Dilute Ammonia
Hibiscus Ixora Pride of Barbados Frangipani Any other flowers'					

The universal indicator

In Experiment 7.2, we observed that acids such as dilute hydrogen chloride acid (hydrochloric acid) and dilute tetraoxosulphate(VI) acid give bright red colours with universal indicator, whereas equal molar concenttrations of some other acids like ethanoic acid and citric acid give a pink (weakly red) colour. Acids of the first type are the strong acids while those of the second type are the weak acids.

Acids vary in degree of acidity from very strong, moderately strong, moderately weak to very weak. There is a similar variation in the degree of basicity of bases from very strong, moderately strong, moderately weak to very weak bases. Basic strength depends on the concentration of hydroxide ions in solution.

The universal indicator is useful for comparing the strength of solutions of acids and bases. It is a mixture of dyes, and shows a wide range of colours like: red, yellow, green, blue, indigo and violet in acid and alkaline solutions. The particular colour assumed in a given solution depends on the degree of acidity or alkalinity of the solution.

Experiment 7.17 Determination of Degrees of Acidity and Basicity of Solutions, Using Universal Indicator
(a) Pour $10cm^3$ 1 mol dm^{-3} hydrogen chloride acid (hydrochloric acid) into a $100cm^3$ measuring cylinder, add distilled water to it up to the $100cm^3$ mark and mix well. The solution you have prepared is 0.1 mol dm^{-3} in strength. Pour $10cm^3$ of it into a labelled test tube.

Solution 2: Pour $10cm^3$ of the remaining 0.1 mol dm^{-3} hydrogen chloride acid (hydrochloric acid) solution into a $100cm^3$ measuring cylinder and add distilled water to it up to the $100cm^3$ mark. Mix well to obtain a solution that is 0.01 mol dm^{-3} Pour $10cm^3$ of it into another test tube.

Solution 3: Take 10cm³ of solution 2 and similarly dilute ten times to obtain a 0.001 mol dm⁻³ solution.

Solutions 4–7 are prepared as above by progressively diluting each preceding solution ten-fold.
    The procedure is summarized in Table 7.9 and Figure 7.20.

TABLE 7.9 Scheme for preparation of a range of molar concentrations of hydrogen chloride acid (hydrochloric acid) solutions.

| | Solution Numbers | | | | | | |
	1	2	3	4	5	6	7
Sample of HCl taken	1.0 mol dm⁻³	Soln. 1	Soln. 2	Soln. 3	Soln. 4	Soln. 5	Soln. 6
Sample volume taken	10.0cm³	10.0cm³	10.0cm³	10.0cm³	10.0cm³	10.0cm³	10.0cm³
Volume of water added	90.0cm³	90.0cm³	90.0cm³	90.0cm³	90.0cm³	90.0cm³	90.0cm³
Total volume	100.0cm³	100.0cm³	100.0cm³	100.0cm³	100.0cm³	100.0cm³	100.0cm³
Volume of aliquot in test tube	10.0cm³	10.0cm³	10.0cm³	10.0cm³	10.0cm³	10.0cm³	10.0cm³

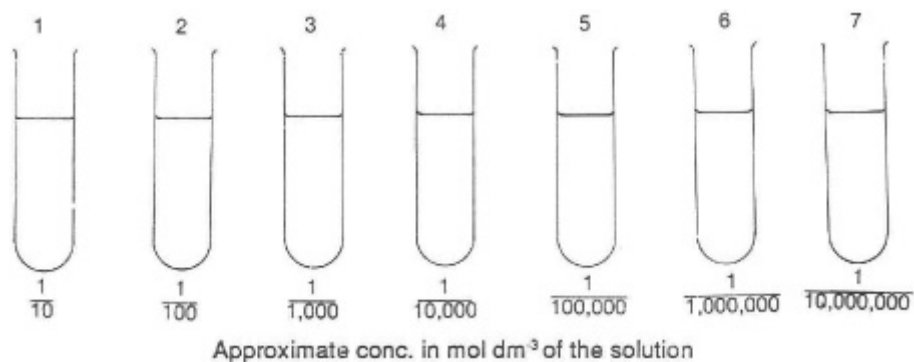| 1 | 2 | 3 | 4 | 5 | 6 | 7 |
| $\frac{1}{10}$ | $\frac{1}{100}$ | $\frac{1}{1,000}$ | $\frac{1}{10,000}$ | $\frac{1}{100,000}$ | $\frac{1}{1,000,000}$ | $\frac{1}{10,000,000}$ |

Approximate conc. in mol dm⁻³ of the solution

Figure 7.20 Resultant concentration in mol dm⁻³

    Place two drops of universal indicator into each solution. Mix the contents thoroughly and observe the colours of the solutions.
(b) Draw up a table of results as shown in Table 7.10.

TABLE 7.10

Volume of 0.1 mol dm⁻³ Alkali Added to 20cm³ 0.1 mol dm⁻³ Ethanoic Acid	0.0 cm³	5.0 cm³	10.0 cm³	15.0 cm³	20.0 cm³	25.0 cm³	30.0 cm³	35.0 cm³
Colour of solution after mixing								

    Then set up a burette containing 35.0cm³ of 0.1 molar sodium

hydroxide. Transfer 20.0cm³ of 0.1 molar ethanoic acid (CH₃COOH) into a conical flask, using a pipette. Add two drops of universal indicator. Shake, and record the colour of the solution. Add the alkali from the burette into the conical flask (Figure. 7.21) in incremental portions of 5.0cm³. Shake the solution and record the colour after each addition until a total of 35.0cm³ of the alkali has been added.
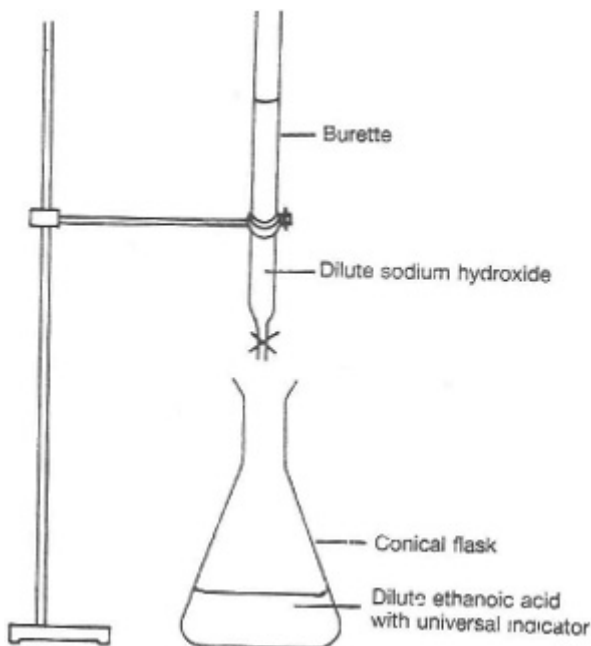

Figure 7.21 Addition of alkali to acid

Draw all the inferences you can about the variation in the degree of acidity and degree of alkalinity in Experiment 7.17 (a) and (b).

The pH Index

The degree of acidity or basicity of a given solution is expressed quantitatively by the pH scale.

The pH scale is a logarithmic expression of the hydrogen ion concentration of a solution. The scale ranges between values of 1 – 14. A solution which is acidic has a pH value below 7, while a solution which is alkaline (basic) has a pH value greater than 7. A neutral solution such as pure distilled water has a pH value of 7. The lower the pH value of a given solution, the more strongly acidic is the solution, while the higher the pH value of a solution, the more strongly alkaline (basic) is the solution.

pH is defined by the expression:

$$pH = -\log_{10} \text{ of molar concentration of hydroxonium ion.}$$
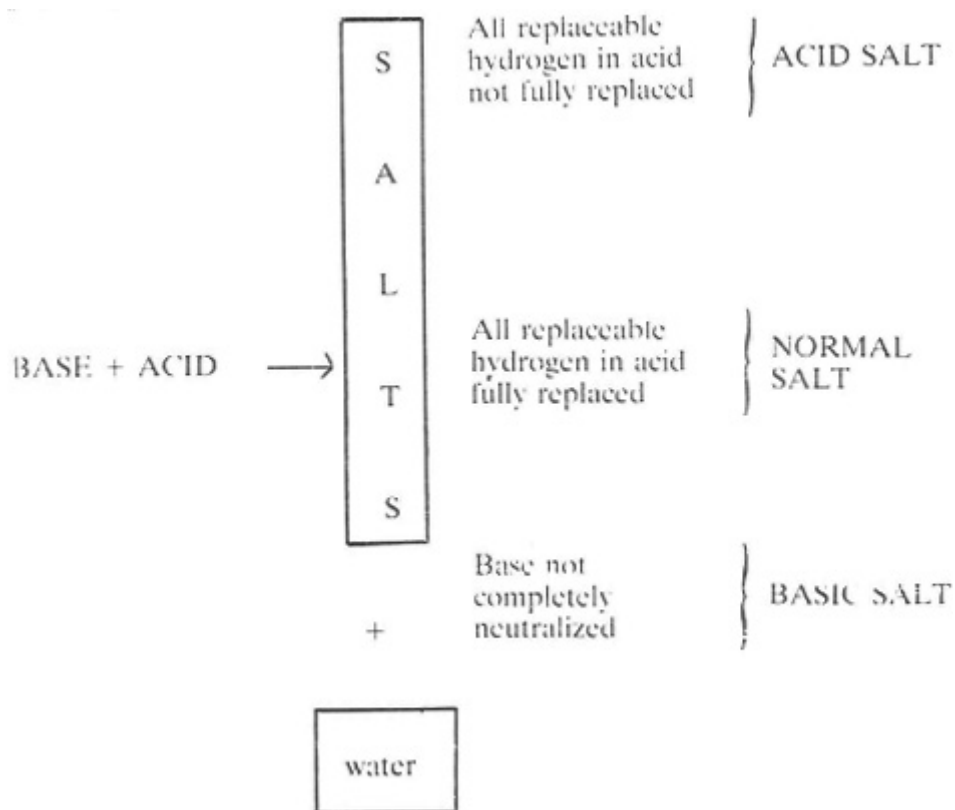$$\text{or } pH = -\log_{10}[H_3O^+]$$

The relationships between the pH of a solution and its degree of acidity or basicity are shown in Table 7.11.

TABLE 7.11 The pH Scale

pH of Solution	Colour of Universal Indicator	Degree of Acidity or Alkalinity
1	} red	} strongly acidic
2		
3	pink	
4	brown	
5	} yellow	} weakly acidic
6		
7	} green	} neutral
8		
9	} blue	} weakly alkaline
10		
11	} indigo	
12		} strongly alkaline
13	} violet	
14		

7.12   Salts

A salt is the substance obtained when the hydrogen ion of an acid is replaced by a metallic ion or the ammonium radical.

BASE + ACID ⟶ S A L T S

All replaceable hydrogen in acid not fully replaced } ACID SALT

All replaceable hydrogen in acid fully replaced } NORMAL SALT

+

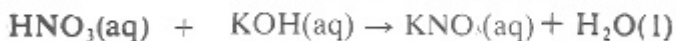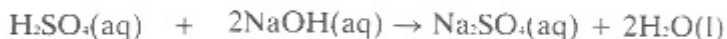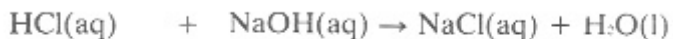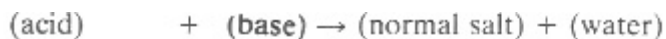Base not completely neutralized } BASIC SALT

water

Normal salts

A normal salt is formed when all the replaceable hydrogen atoms of an

acid have been replaced by metal.

Examples

(acid)   +   (base) → (normal salt) + (water)

$HCl(aq)$ + $NaOH(aq) → NaCl(aq) + H_2O(l)$

$H_2SO_4(aq)$ + $2NaOH(aq) → Na_2SO_4(aq) + 2H_2O(l)$

$HNO_3(aq)$ + $KOH(aq) → KNO_3(aq) + H_2O(l)$

$H_3PO_4(aq)$ + $3NaOH(aq) → Na_3PO_4(aq) + 3H_2O(l)$
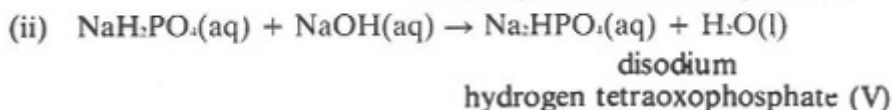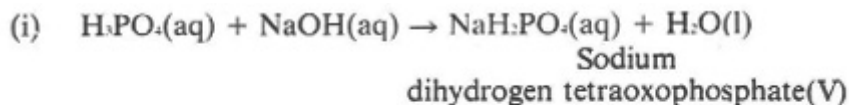
Acid salts

We have learnt that acids with basicity greater than one, such as dibasic and tribasic acids, ionise in steps, (i.e. the replaceable hydrogen atoms go into solution one after the other). With tetraoxosulphate(VI) acid for example, the first step of neutralization is:

$$H_2SO_4(aq) + NaOH(aq) → NaHSO_4(aq) + H_2O(l)$$
Sodium
hydrogen tetraoxosulphate(VI)
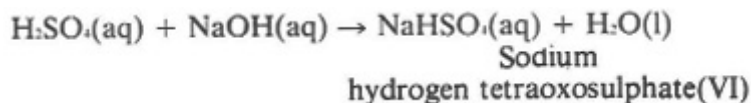
The product is an acid salt because it still contains a replaceable hydrogen atom.

Other examples of the formation of acid salts are:

(i) $H_3PO_4(aq) + NaOH(aq) → NaH_2PO_4(aq) + H_2O(l)$
Sodium
dihydrogen tetraoxophosphate(V)

(ii) $NaH_2PO_4(aq) + NaOH(aq) → Na_2HPO_4(aq) + H_2O(l)$
disodium
hydrogen tetraoxophosphate (V)

Basic salts

Bases formed from di- or tri-valent metals may be neutralized by acids in steps. When such a base is incompletely neutralized by an acid, the salt formed is a basic salt

Examples

$$Zn(OH)_2(aq) + HCl(aq) → Zn(OH)Cl(aq) + H_2O(l)$$
Basic zinc chloride
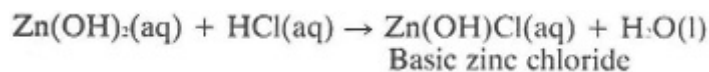
Other basic salts include: basic lead trioxocarbonate(IV), $Pb(OH)_2.2PbCO_3$; basic magnesium chloride, $Mg(OH)Cl$.

TABLE 7.12 Types of Salts and How They Are Formed

Normal salts	Formed by complete neutralization of acid and base
Acid salt	Formed by incomplete neutralization of acid
Basic salt	Formed by incomplete neutralization of base

7.13   Properties of salts

(a) **Solubility in water:** Salts differ in their ability to dissolve in water.

Some salts dissolve readily in water whereas others are insoluble or only sparingly soluble.

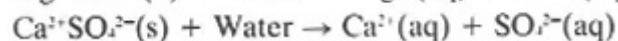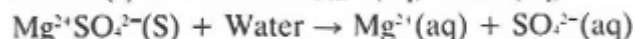The general solubility patterns among salts are shown in Table 7.13

TABLE 7.13 Solubility of Salts

Type of Salt	Solubility Characteristics
* Trioxocarbonates(IV) * Sulphides * Trioxosulphates(IV)	All are insoluble except the potassium, sodium and ammonium salts
Chlorides	All are soluble except silver and lead (II) chlorides (Note: lead(II) chloride is soluble in hot water)
Trioxonitrates(IV)	All are soluble
* Tetraoxosulphates(VI)	All soluble except barium and lead tetraoxosulphate(VI).CaSO$_4$ is sparingly soluble

*** Note that the acid salts are generally soluble.**

(b) **Ionisation in water:** All salts exist as ions even in the solid state.

All salts once dissolved in water are completely ionised into mobile ions.

Examples

$Na^+Cl^-(s) + Water \rightarrow Na^+(aq) + Cl^-(aq)$
$Mg^{2+}SO_4^{2-}(S) + Water \rightarrow Mg^{2+}(aq) + SO_4^{2-}(aq)$
$Ca^{2+}SO_4^{2-}(s) + Water \rightarrow Ca^{2+}(aq) + SO_4^{2-}(aq)$

Note that even though calcium tetraoxosulphate(VI) dissolves only slightly in water, the portion which dissolves will ionise completely in water.
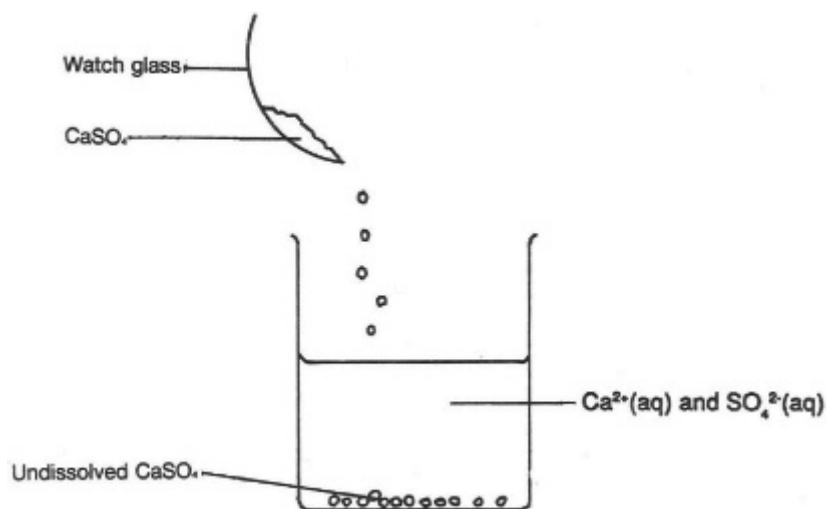
Figure 7.22 Ionisation of CaSO₄ in water

(c) **Hydration of salts:** Salts are often obtained in the solid state by crystallisation from their solutions in water. The solid state of the salts may assume one of two forms; either an amorphous powder or a regular crystalline structure. Most salts incorporate a fixed proportion of molecules of water in the process of forming their crystalline structure.
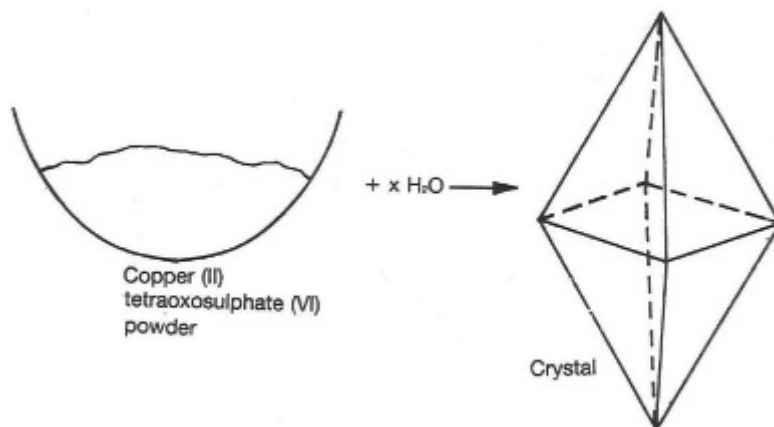

Figure 7.23 Crystal formation in a salt solution

This water is called **water of crystallization.** Any crystalline salt containing water of crystallization is called a **hydrated salt** or simply a **hydrate.** Salts without water of crystallization are called **anhydrous** salts. They are often powdery in nature.

TABLE 7.14 Examples of Hydrated Salts

Formula of Salt Hydrate	Common Name	IUPAC Name
$Na_2SO_4 \cdot 10H_2O$		Sodium tetraoxosulphate (VI) decahydrate
$CaCl_2 \cdot H_2O$	—	Calcium chloride monohydrate
$CaCl_2 \cdot 2H_2O$	Gypsum	Calcium chloride dihydrate
$CuSO_4 \cdot 5H_2O$	Blue Vitriol	Copper(II) tetraoxosulphate(VI) pentahydrate
$FeSO_4 \cdot 7H_2O$	Green Vitriol	Iron(II) tetraoxosulphate(VI) heptahydrate
$MgSO_4 \cdot 7H_2O$	Epsom salt	Magnesium tetraoxosulphate(VI) heptahydrate
$Na_2CO_3 \cdot 10H_2O$	Washing soda	Sodium trioxocarbonate(IV) decahydrate

Experiment 7.18 Investigating The Nature of Hydration

Accurately weigh a hard glass test-tube. Put about 5.0g of hydrated copper(II) tetraoxosulphate(VI) crystals into the test-tube. Weigh the test tube again to determine the accurate mass of the salt. Record the colour and texture of the crystals.

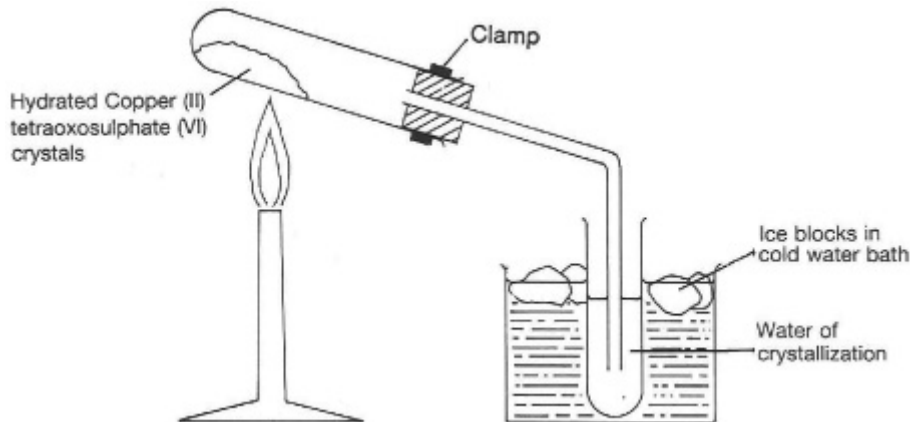Stopper the test tube with a cork which carries a bent delivery tube as illustrated in Figure 7.24

Figure 7.24 Removal of water of crystallization by heating

Allow the other end of the delivery tube to dip into a second weighed test-tube which is placed in a bath of cold water as shown. Heat the test-tube containing the crystals, gently at first, then strongly. Continue the heating until no further liquid collects in the other test tube. Allow the test tube with the residue to cool down in a desiccator. Then weigh the test-tube again to determine the mass of the residue. Heat the test-tube again for 10 minutes, cool and weigh again until the mass of test-tube and residue remain constant for two successive weighings. Record the final mass, and hence determine the loss of

mass of the crystals.

Finally, weigh the second test-tube containing the condensed liquid and determine the mass of liquid collected.

(a) Carry out the following tests on the residue:

(i) Record its colour and texture.

(ii) Put some of the residue in a test-tube. Add some drops of water. What is the new colour of the sample?

(b) Test the liquid collected in the other tube with dry cobalt(II) chloride paper. Identify the liquid.

From your experimental results calculate,

(i) The mass of water lost by the hydrate (compare this with the mass of liquid collected).

(ii) The number of moles of water contained per mole of $CaSO_4$ in the crystal.

Hence, determine the value of x in the formula $CuSO_4xH_2O$ for the hydrate.

Write the equation for the change which occurs on heating the hydrated salt.

Specimen calculation

Mass of test-tube + hydrate = b g
Mass of test-tube + residue = c g
Mass of empty test-tube   = a g.

∴ **Mass** of hydrate   = (b – a) g
∴ Mass of residue   = (c – a) g

Therefore mass of water of hydration lost

$$= (b – a)–(c – a)g$$
$$= (b – c)g$$

(c – a)g of $CuSO_4$ (anhydrous) was hydrated by
(b – c) g of water.
Molar mass of $CuSO_4$ = 159.5 g
Molar mass of water = 18 g

Thus 159.5 g of $CuSO_4$ will be hydrated by

$$\frac{(b – c) \times 159.5}{(c – a)} \text{ g of water}$$

$$= \frac{(b – c) \times 159.5}{(c – a)} \times \frac{\text{moles of water}}{18}$$

$$\therefore x = \frac{(b – c) \times 159.5}{(c – a) \times} \times \frac{\text{moles of water}}{18}$$

All hydrated copper(II) salts are blue. It follows that the blue colour is due to the hydration of the copper(II) ion and not the acid radical in the salt.

Generally it is the metal ion, not the acid radical, that is hydrated in

the crystal.

Thus: $Cu^{2+}(s) + xH_2O(l) \rightarrow [Cu(H_2O)x]^{2+}(aq)$

     (colourless)   (blue)
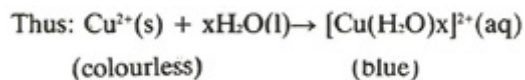
Double Salts

A double salt is a compound made up of two salts that crystallize from solution in definite proportions by mass. An example is potassium alum (or potash alum) of formula $K_2SO_4.Al_2(SO_4)_3.24H_2O$

In this compound, one molecule of potassium tetraoxosulphate(VI), one molecule of aluminium tetraoxosulphate(VI) and twenty-four molecules of water crystallize from solution to form a giant molecule of the alum.

$K_2SO_4(aq) + Al_2(SO_4)_3(aq) + 24H_2O(l) \rightarrow K_2SO_4. Al_2(SO_4)_3.24H_2O(aq)$
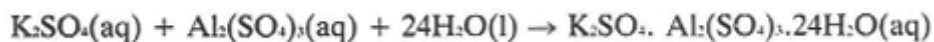
The crystals of potassium alum have eight faces, each forming a triangle.

When dissolved in water, the double salt behaves just as a simple mixture consisting of K^+, Al^{3+} and SO_4^{2-} ions.

By replacing the K^+ with Na^+, Li^+ or NH_4^+ and/or the Al^{3+} with Cr^{3+} or Fe^{3+} we obtain an important group of double salts generally called alums.

Alums have the general formula,

$M^+_2(SO_4^{2-}).M_2^{3+}(SO_4^{2-})_3.24H_2O$
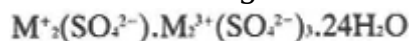
where M^+ is a monovalent metal, and M^{3+} is a trivalent metal.

Note that not all double salts are alums. One double salt which is not an alum is $MgCl_2.KCl.6H_2O$.

# 7.14 Methods for the preparation of salts

1. By neutralization and crystallization

Experiment 7.19 Preparation of a normal salt (sodium tetraoxosulphate(VI) decahydrate)

Pour about 40cm³ of 0.05 mol dm⁻³ tetraoxosulphate(VI) acid into a burette. Record the initial level in the burette. Then using a pipette measure accurately 25.0cm³ of 0.1 mol dm⁻³ sodium hydroxide solution into a conical flask. Add a few drops of methyl orange into the flask.

Cautiously run the acid from the burette into the alkali solution while swirling the flask to stir the solution. After adding up to 24 cm³ of the acid, continue the addition drop by drop, swirling after each addition, until the colour just changes.

Record the final burette reading. Hence, by substraction of the final burette readings, determine the volume of acid which neutralizes the 25 cm³ of the sodium hydroxide.

$$2NaOH(aq) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$$
$$\text{(normal salt)}$$
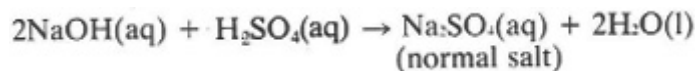
We now know the quantities of the acid and the base required for preparing the normal salt.

Now discard the contents of the conical flask. Using a clean flask, pipette 25 cm³ of the given alkali then add from a burrete the exact volume of the acid determined above. Shake to mix thoroughly.

To obtain crystals of the salt, pour the salt solution into an evaporating dish. Heat it on a wire guaze to concentrate it until a test sample on a glass rod gives crystals on cooling. Then allow the evaporating dish to cool to obtain crystals. Filter off and collect the crystals of $Na_2SO_4 . 1OH_2O$.

Experiment 7.20 Preparation of an acid salt (sodium hydrogen tetraoxosulphate(VI))

Determine (as in Experiment 7.19) the volume of 0.05 mol dm⁻³ tetraoxosulphate(VI) which neutralizes 25.0 cm³ of the given 0.1 mol dm⁻³ sodium hydroxide solution.

To prepare the acid salt, pipette 25.0 cm³ of the alkali into a comcal flask as before, then add from a burette exactly twice the volume of the acid given by the titration. Mix the solutions throughly and then crystallize the salt formed as in Experiment 7.19. Filter and collect the crystals. The product consists of crystals of sodium hydrogen tetraoxosulphate(VI).

$$NaOH(aq) + H_2SO_4(aq) \rightarrow NaHSO_4(aq) + H_2O(l)$$
$$\text{(acid salt)}$$
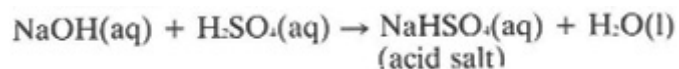
Note that the amount of sodium hydroxide needed to prepare sodium hydrogen tetraoxosulphate(VI) is half the amount used in preparing sodium tetraoxosulphate(VI) the normal salt.

Other salts which may be prepared by neutralization include ammonium .trioxonitrate(V), sodium chloride, sodium ethanedioate, sodium ethanoate and potassium dihydrogen tetraoxophosphate(V).
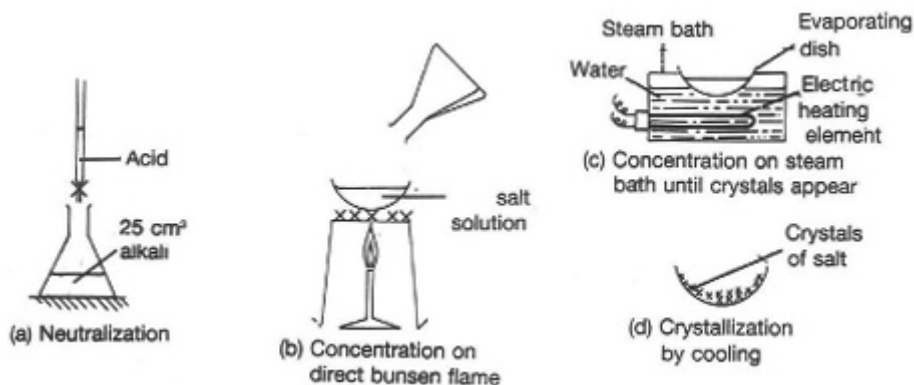
Figure 7.25 Preparation of a salt by neutralization and crystallization

2. By the reaction between a gaseous acid anhydride and an alkali

Experiment 7.21 Preparation of Sodium Hydrogen Trioxosulphate(VI) from Sulphur(IV) Oxide

Set up in a fume cupboard the apparatus shown in Figure 7.26. Pipette 25.0 cm³ of concentrated sodium hydroxide solution, into a beaker, and saturate the solution with sulphur(IV) oxide produced from the set-up in Figure 7.26.
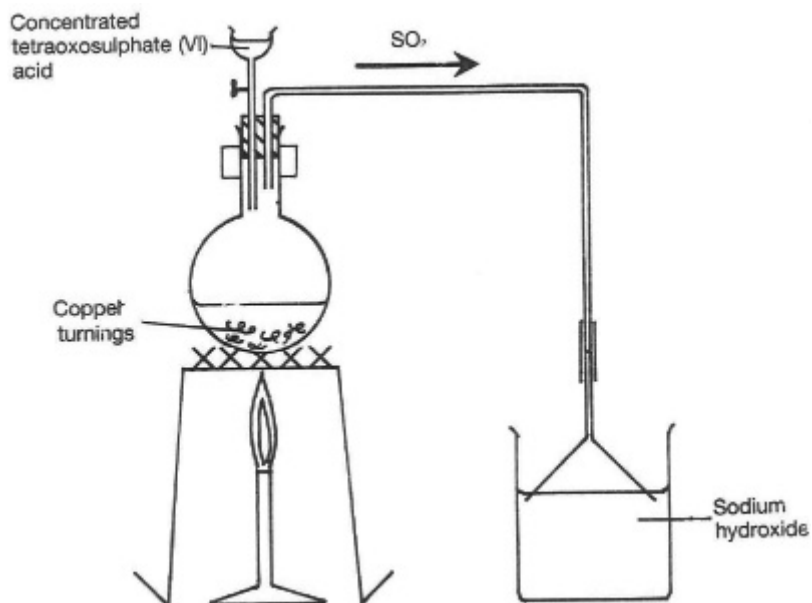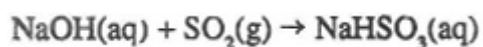

**Figure 7.26 Preparation of NaHSO₃ and Na₂SO₃**

Allow the sulphur(IV) oxide to pass into the alkali solution until the vapour over the alkali in the beaker is strongly acidic to litmus, then remove the beaker. Crystallize the salt formed.

$$NaOH(aq) + SO_2(g) \rightarrow NaHSO_3(aq)$$

Experiment 7.22 Preparation of sodium trioxosulphate(IV) crystals from sulphur(IV) oxide

Pipette 25.0 cm³ aliquots of concentrated sodium hydroxide solution into two separate beakers.

Saturate the solution in one beaker with sulphur(IV) oxide as described in Experiment 7.21, then add the other 25.0 cm³ portion of sodium hydroxide into this beaker, while stirring. Crystallize the salt formed as before.

$$NaHSO_3(aq) + NaOH(aq) \rightarrow Na_2SO_3(aq) + H_2O(l)$$
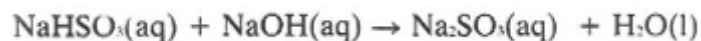
Some normal and acid salts which could be prepared by the action of the acid anhydride on alkali are indicated below in Table 7.15.

TABLE 7.15

Gaseous Anhydride	Alkali	Normal salt	Acid salt
CO_2	NaOH	Na_2CO_3	$NaHCO_3$
SO_2	NaOH	Na_2SO_3	$NaHSO_3$
H_2S	NaOH	Na_2S	NaHS
SO_2	KOH	K_2SO_3	$KHSO_3$
CO_2	$Ca(OH)_2$	$CaCO_3$	$Ca(HCO_3)_2$
SO_3	$Ca(OH)_2$	$CaSO_4$	$Ca(HSO_4)_2$

3. By the action of acids on metals

Acids react with metals in two ways:

(i) Displacement of hydrogen by light metals.

Example

$$Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$$

(ii) Oxidation of the metal by an oxidizing acid.

Example:

$$3Cu(s) + 8HNO_3(aq) \rightarrow 3Cu(NO_3)_2(aq) + 2NO(g) + 4H_2O(l)$$

Experiment 7.23 Preparation of Magnesium Tetraoxosulphate(VI) Crystals

Pour dilute tetraoxosulphate(VI) acid into a beaker. Add pieces of magnesium ribbon into the acid until effervescence stops. Add more magnesium ribbon until it does not react further. Filter off the excess magnesium, then crystallize the solution. Crystals of magnesium tetraoxosulphate(VI) heptahydrate and iron(II) tetraoxosulphate(VI) heptahydrate can be prepared in the same way.

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

$MgSO_4(aq) + 7H_2O(l) \rightarrow MgSO_4. 7H_2O(aq)$
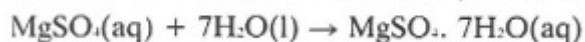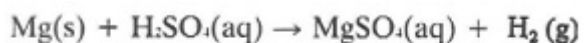
Experiment 7.24 Preparation of copper(II) trioxonitrate(V) Crystals
Pour about 10 cm³ of concentrated trioxonitrate(V) acid into a beaker. Heat the acid gently in a fume cupboard while adding copper turnings until no more copper dissolves.

   Continue the heating until the brown gas (nitrogen(IV) oxide) is completely driven off. Do not evaporate to dryness. Add some water to the contents of the beaker to dissolve the residue. Filter off any unreacted copper left. Then crystallize the salt as before.

   The greenish-blue crystals are $Cu(NO_3)_2. 6H_2O$.

$Cu(s) + 4HNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + 2NO_2(g) + 2H_2O(l)$

4. By the reaction of an insoluble base with an acid

   Lead(II) trioxonitrate(V) may be prepared by pouring about 10 cm³ dilute trioxonitrate(V) acid into a beaker, and adding lead(II) oxide (litharge) a little at a time until no more dissolves. The undissolved oxide is then filtered off, and the filtrate crystallized to obtain white lead(II) trioxonitrate(V) crystals.
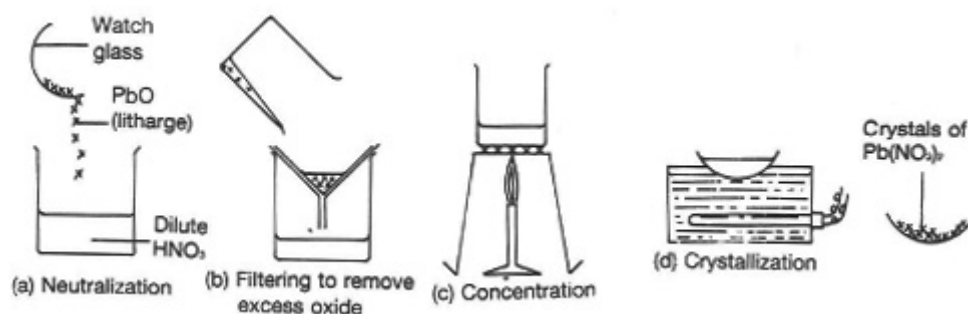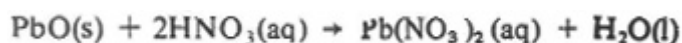
$PbO(s) + 2HNO_3(aq) \rightarrow Pb(NO_3)_2 (aq) + H_2O(l)$

Figure 7.27 Preparation of Pb(NO₃)₂ crystals from PbO

5. By the reactions of trioxocarbonates(IV) with dilute acids

Experiment 7.25 Preparation of copper (II) chloride from copper trioxocarbonate(IV)
Pour about 10cm³ of dilute hydrochloric acid into a beaker, and add copper(II) trioxocarbonate(IV). Warm the mixture to expel the carbon(IV) oxide evolved. Add more copper(II) trioxocarbonate(IV) until excess solid remains.
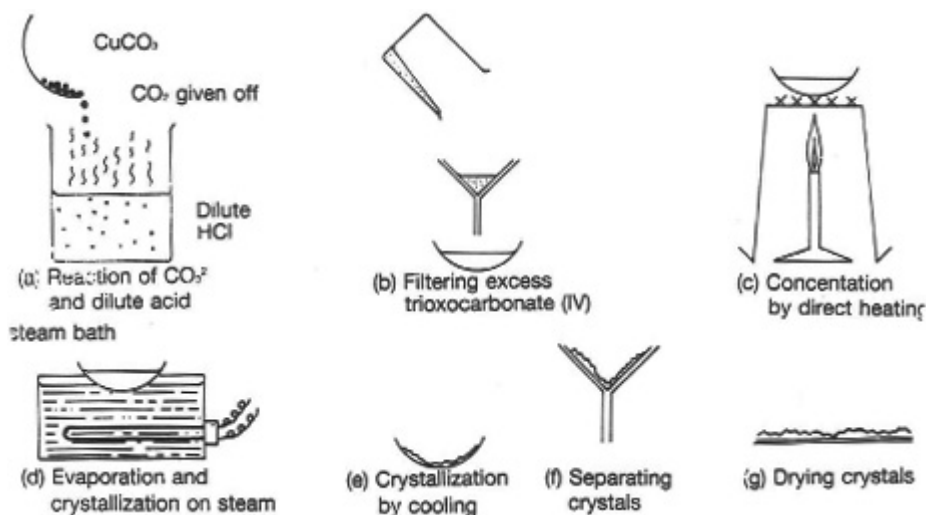
Figure 7.28 Preparation of copper(II) chloride from copper(II) trioxocarbonate (IV)

Filter off the excess solid, concentrate the solution and crystallize it to obtain copper(II) chloride hexahydrate, $CuCl_2.6H_2O$.

$$CuCO_3(s) + 2HCl(aq) \rightarrow CuCl_2(aq) + CO_2(g) + H_2O(l)$$
$$CuCl_2(aq) + 6H_2O(l) \rightarrow CuCl_2.6H_2O(aq)$$

Any trioxocarbonatc(IV) could be used with any dilute acid if the salt formed by the reaction is soluble. However, this method is not suitable for preparing calcium tetraoxosulphate(VI) which is sparingly soluble.

6. By direct combination, of elements

In the method of direct combination, two elements react together to form a salt. The method is useful for preparing chlorides, bromides and sulphides, i.e. binary compounds.

Experiment 7.26 Preparation of Iron(III) Chloride
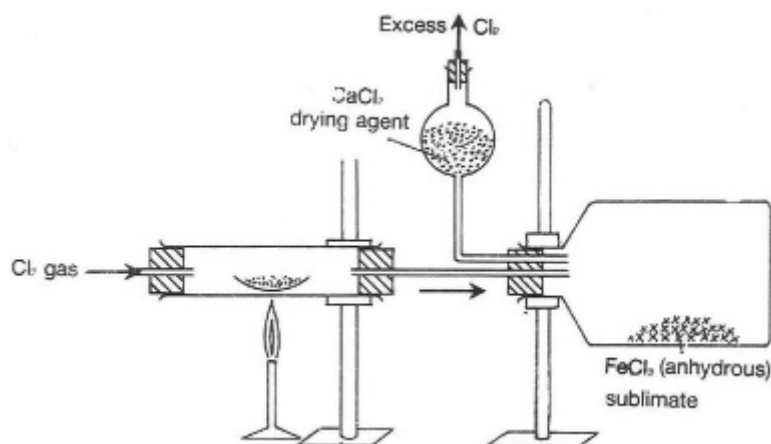Set up the apparatus as shown in Figure 7.29

Figure 7.29 Preparation of Iron(III) chloride

Put about 5 g of fresh and clean iron filings in a boat and slip the boat into the combustion tube which is clamped horizontally as shown. The boat is heated while chlorine gas is passed over the iron filings in the combustion tube. The hot iron filings combine with the chlorine to produce iron(III) chloride. This salt sublimes and its vapour is passed into the receiving jar where it changes to a solid.

Iron(III) chloride is readily hydrolysed by water. It is therefore necessary that the anhydrous salt be prevented from contact with water from the atmosphere. This is achieved by allowing any excess chlorine gas to escape from the jar through a drying tube containing fused calcium chloride.

$$2Fe(s) + 3Cl_2(g) \xrightarrow{heat} 2FeCl_3(s)$$
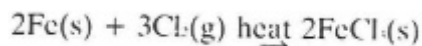
A similar procedure is used for preparing iron(II) chloride, though hydrogen chloride gas rather than chlorine is used in this case.

$$Fe(s) + 2HCl(g) \rightarrow FeCl_2(s) + H_2(g)$$

Experiment 7.27 Preparation of magnesium chloride
Place a piece of magnesium ribbon in a deflagrating spoon and light this at one end. Then lower it into a gas jar filled with chlorine gas. Allow the magnesium ribbon to burn completely, then collect the residue left at the bottom of the jar.

$$Mg(s) + Cl_2(g) \rightarrow MgCl_2(s)$$
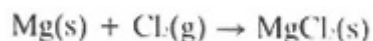
7. Methods for preparing insoluble salts

(a) By double decomposition
In a double decomposition reaction, an insoluble salt and another soluble product are obtained by the exchange of radicals between two compounds both of which are in solution.

Experiment 7.28 Preparation of lead(II) chloride
Pour about 20 cm³ of a dilute solution of lead(II) trioxonitrate(V) into a beaker. Add dilute hydrochloric acid while stirring, until no more precipitation occurs. Allow the mixture to stand for a few minutes. Filter off the precipitate and dry the solid between filter papers (Figure 7.30).

$$Pb(NO_3)_2(aq) + 2HCl(aq) \rightarrow PbCl_2(s) + 2HNO_3(aq)$$
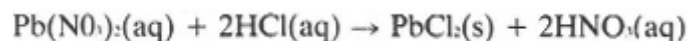
The following double decomposition reactions can be used to prepare other insoluble salts.

$$AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$$
$$Pb(NO_3)_2(aq) + 2NaCl(aq) \rightarrow PbCl_2(s) + 2NaNO_3(aq)$$
$$BaCl_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2NaCl(aq)$$
$$CuSO_4(aq) + Na_2CO_3(aq) \rightarrow CuCO_3(s) + Na_2SO_4(aq)$$
$$CaCl_2(aq) + Na_2SO_4(aq) \rightarrow CaSO_4(s) + 2NaCl(aq)$$
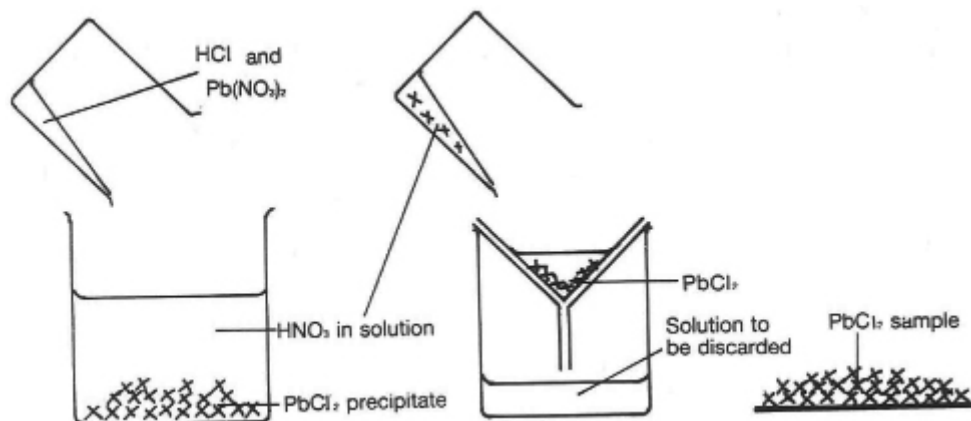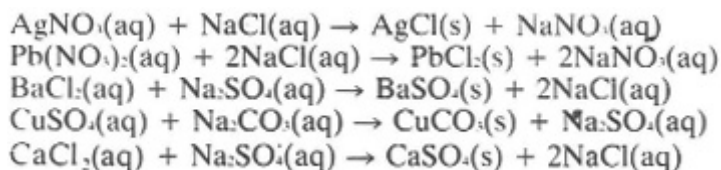

Figure 7.30 Preparation of salt by double decomposition

(b) By direct combination

The method of direct combination used for preparing soluble salts may also be used for preparing insoluble salts.

Experiment 7.29 Preparation of iron sulphide

Place 2.8g of iron filings in a crucible. Add 1.6g of sulphur powder. Mix the solids thoroughly with a glass rod. Heat the crucible over a bunsen flame, gently at first and then more strongly while stirring the mixture (Figure 7.31). Stop heating when the mixture appears uniform.

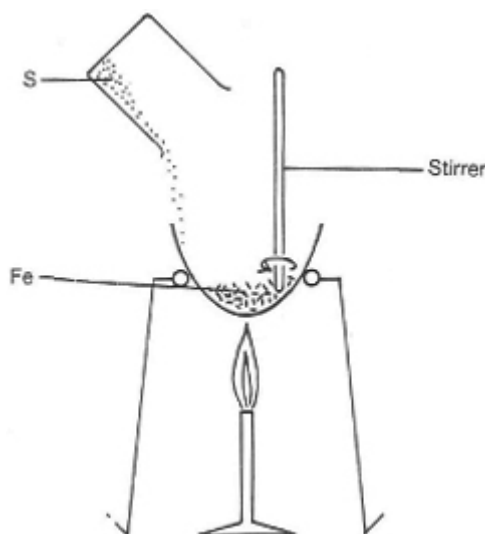The black metallic product is iron(II) sulphide, FeS.

$$Fe(s) + S(s) \rightarrow FeS(s)$$

Figure 7.31 Preparation of iron(II) sulphide by direct combination

7.15 Efflorescence, hygroscopy and deliquescence

If a hydrated salt is exposed to the atmosphere at room temperature and it loses some of its water of crystallization, the salt is said to be **efflorescent.** This action is called **efflorescence.** When this happens the physical properties of the salt such as its crystalline nature and colour may change.

Some other substances on exposure to air absorb water-vapour from the atmosphere. The water absorbed is not enough to form a solution. The substance is said to be **hygroscopic.** But if the substance absorbs enough water-vapour from the air to form a solution it is said to be **deliquescent.**

We shall investigate these properties in some compounds in Experiment 7.30

Experiment 7.30 To identify efflorescent, hygroscopic and deliquescent substances

(i) sodium trioxocarbonate(IV) decahydrate crystals,

(ii) concentrated tetraoxosulphate(VI) acid (about 10cm³),

(iii) calcium oxide,

(iv) magnesium tetraoxosulphate(VI) heptahydrate crystals,

(v) sodium hydroxide pellets,

(vi) calcium chloride,

Weigh each watch glass and contents and record the weights. Place the watch glasses on a side table in the laboratory until your next class. Then weigh each watch glass again. Record any changes in mass and physical appearance.

From the experiment, you will observe that:

(i) The sodium trioxocarbonate(IV) decahydrate crystals turned into a white powder and lost weight. It has lost some of its water of crystallization.

$$Na_2CO_3.10H_2O(s) \rightarrow Na_2CO_3.H_2O(s) + 9H_2O(g)$$
(glassy crystals)   (white powder)
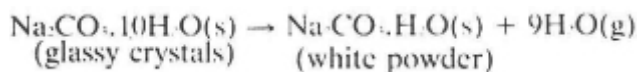
Sodium trioxocarbonate(IV) decahydrate is therefore an efflorescent compound.

Other efflorescent compounds include magnesium tetraoxosulphate(VI) heptahydrate, and sodium heptaoxoborate(III) decahydrate crystals.

Na_2B_4O-. $10H_2O$ (borax).

(ii) Though concentrated tetraoxosulphate(VI) acid and calcium oxide, have not changed physically, they have gained weight by absorbing some water vapour from the air. They are therefore hygroscopic substances. Other hygroscopic compounds are copper(II) oxide, and sodium trioxonitrate(V).

(iii) The sodium hydroxide pellets and solid calcium chloride absorbed enough water-vapour from the air and formed solution. They are deliquescent substances. Other deliquescent substances are iron(III) chloride and calcium trioxonitratc(V).

Because of the affinities of hygroscopic and deliquescent substances for water, they are often used as drying agents. Concentrated tetraoxosulphate(VI) acid and calcium chloride are common drying agents for neutral and acidic gases. Calcium oxide is used to dry ammonia because ammonia reacts with concentrated tetraoxosulphate(VI) acid and calcium chloride.

Based on your observations in these experiments, the substances investigated above can be classified into:

(1) Efflorescent compounds e.g. $MgSO_4.7H_2O$ and Na_2CO_3. $10H_2O$, which lose weight on exposure to air due to loss of water of crystallization.

(2) Hygroscopic compounds e.g. calcium oxide and concentrated tetraoxosulphate(VI) acid which gain weight because they have absorbed water from the air but they do not dissolve in the water.

(3) Deliquescent compounds e.g. sodium hydroxide pellets and calcium chloride, which gain in weight because they absorb water and dissolve in the water.
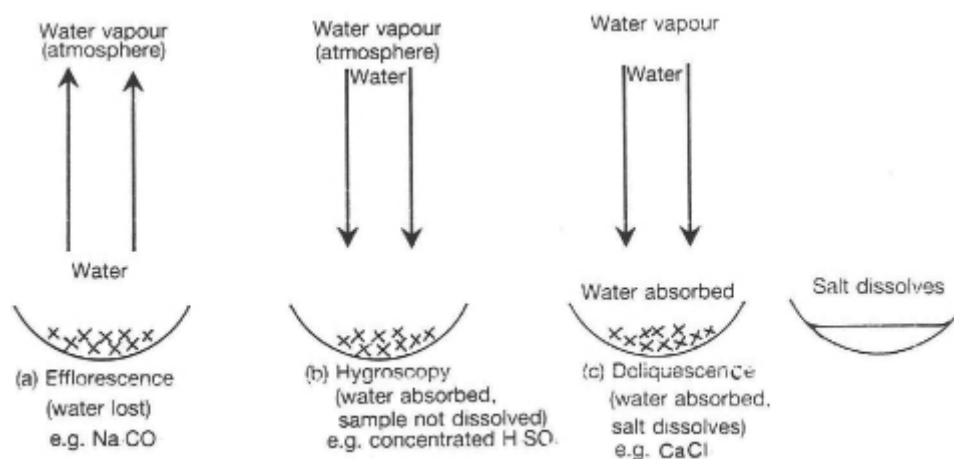

Figure 7.32 Efflorescence, Hygroscopy, and Deliquescence

Uses of Hygroscopic and Deliquescent Substances

Concentrated tetraoxosulphate(VI) acid is used for drying most gases, e.g. hydrogen chloride, chlorine and carbon(IV) oxide.

Fused calcium chloride is used as a desiccant in desiccators, in U-tubes for drying gases: carbon(IV) oxide, oxygen, hydrogen and for drying most organic liquids. Calcium oxide is used for drying ammonia gas.
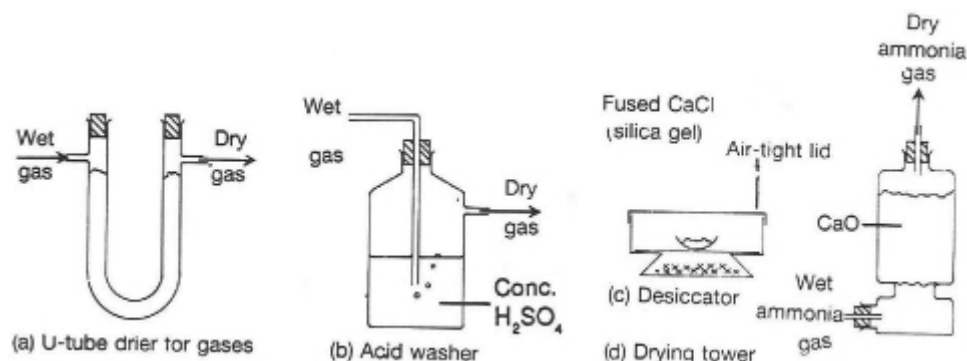

Figure 7.33 Types of drying apparatus and drying agents

The hydrate, gypsum, calcium tetraoxosulphate(VI) dihydrate, $CaSO_4$, $2H_2O$, has very interesting properties. On heating, it loses part of its water of crystallization to form Plaster of Paris ($CaSO_4$) H_2O. When the 'plaster of Paris' (P.O.P) is mixed with water, it sets in a few minutes to reform the hard crystalline gypsum by recovering the water lost. During setting, the material expands slightly forming a tight casing or mould. Gypsum is therefore extensively used for making surgical casts.
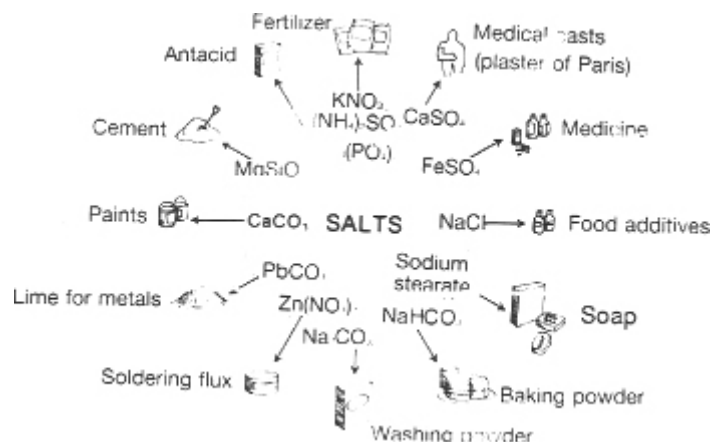

Figure 7.34 Uses of Salts

Chapter Summary

Acids

Characteristic properties of acids:

(i)   Change colour of indicators e.g. change litmus from blue to red.

(ii)   React with metals to give hydrogen and a salt.

(iii)   Liberate carbon(IV) oxide with trioxocarbonates(IV) and hydrogen trioxocarbonates(IV).

(iv)   Neutralize bases.

(v)   Have sour tastes.(CAUTION: Do not taste chemicals)

An acid is a substance which produce hydroxnium ions in aqueous solution.

Acids are prepared by :

(i)   Reaction between an acidie anhydride and water.

(ii)   Displacement of weak and/or more volatile acids from their salts by a stronger and/or less volatile acid

(iii)   Double decomposition reactions involving the precipitation of insoluble salts, leaving an acid in solution.

Bases

Characteristic properties of bases:

(i)   Change colour of indicators, e.g. change litmus from red to blue.

(ii)   Neutralize acids.

(iii)   Have a soapy slippery feel.

(iv)   Have bitter tastes.(CAUTION: Do not taste chemicals)

A base yields hydroxide ions in aqueous solution. Bases are prepared by:

(i)   The reaction between a metal and oxygen.

(ii)   The reaction between light metals and water (steam).

(iii)   The reaction between a metallic oxide and water.

(iv)   The decomposition of oxygen-containing salts such as lead(II) trioxonitrate(V)

(v)   Precipitation of insoluble bases by alkalis.

(vi)   Displacement of a volatile base (e.g. ammonia) from its salts by caustic alkali.

Soluble bases are called alkalis. Examples are sodium hydroxide (caustic soda), potassium hydroxide (caustic potash).

Neutralization occurs between the hydroxonium (H_2O) ion from an acid and the hydroxide (OH) ion from a base to form water.

Acid-base indicators are dyes which change colour depending on the concentration of H_2O and OH ions in the solution.

Degree of acidity/basicity is a measure of the ratio of concentration of H_2O and OH in solution.

it can be determined by:

(1) the colour of universal indicator paper or solution.

(2) use of a pH meter

Salts

Salts are formed from acids by the replacement of one or more hydrogen atoms with metals.
Salts may be (a) soluble or insoluble
                    (b) hydrated or anhydrous.

Types of salts

(i)   An acid with more than one replaceable hydrogen can form normal and acid salts.
(ii)   A base with more than one hydroxide group can form normal and basic salts.
(iii) Double salts are formed when two salts crystallize from the same solution in definite proportions, e.g. $K_2SO_4.Al_2(SO_4)_3.24H_2O$ (potassium alum)

Preparation of salts

Soluble salts are prepared by:
(i)   Neutralization and crystallization
(ii)   The reaction between a gaseous acid anhydride and an alkali
(iii) The action of acids on metals
(iv)   The action of acids on metal oxides
(v)   The action of acids on trioxocarbonates(IV) or hydrogen trioxocarbonates(IV)
(vi)   Direct combination of elements
Insoluble salts are prepared by:
(i)   Double decomposition
(ii)   Direct combination of elements

Efflorescence, hygroscopy and deliquescence.

Efflorescence is a phenomenon whereby certain hydrated salts lose their water of crystallization on exposure to the atmosphere.
Hygroscopy is a phenomenon whereby certain substances absorb moisture on exposure to the atmosphere, but do not dissolve in the moisture.
Deliquescence is a phenomenon whereby certain substances absorb moisture on exposure to the atmosphere and dissolve in it.

Assessment

1. (a) Describe two methods for the laboratory preparation of hydrogen chloride acid (hydrochloric acid).
   (b) Describe what you would observe if the dilute acid

    (i)   is reacted with calcium trioxocarbonate(IV).

    (ii)   is added gradually to a solution of sodium hydroxide to which few drops of phenolphthalein indicator have been added.

    (iii)   is brought in contact with blue and red litmus papers.

    (iv)   is treated with a piece of magnesium ribbon.

    (v)   is heated.

(c) Write balanced equations for the reactions in (b) where possible.

2. (a) Define the terms — acid, base, neutralisation, acid salt, basic salt and basicity of an acid.

  (b)   List five common mineral acids and indicate their basicities.

  (c)   Describe how you would determine the heat of neutralisation of a strong acid by a strong base.

  (d)   List six common bases, and their commonest applications.

3. (a) What are acid salts, basic salts, hygroscopic compounds, and efflorescent compounds?

  (b)   Describe in detail how you would prepare pure potassium chloride by the action of an acid with a base.

  (c)   Would the solutions of the following salts in water be acidic, basic or neutral: calcium chloride, iron(II) tetraoxosulphate(VI), potassium trioxocarbonate(IV), sodium ethanoate, ammonium chloride and magnesium trioxonitrate(V)?

    Give reasons

4. (a) List four acid-base indicators, and their respective colours in acidic and alkaline solutions.

  (b)   How would you compare the strengths of different solutions of an acid and an alkali using a universal indicator solution?

  (c)   What volume of 0.05 mol dm^{-3} tetraoxosulphate(VI) acid should be added to 50cm^3 of 0.1 mol dm^{-3} potassium hydroxide solution in order to prepare only a pure acid salt?

    Write an equation for the reaction.

5. Write equations for the preparation of the following salts by the stated methods.

    (i)   Potassium trioxonitrate(V) by neutralisation.

    (ii)   Sodium hydrogen tetraoxosulphate(VI) by neutralisation.

    (iii)   Magnesium chloride from calcium trioxocarbonate(IV).

    (v)   Iron(II) sulphide by direct combination of elements.

    (vi)   Barium tetraoxosulphate(VI) by double decomposition.

    (vii)   Copper(II) trioxocarbonate(IV) from copper metal and trioxonitrate(V) acid.

6. Why

  (a)   is concentrated tetraoxosulphate(VI) acid not used for drying ammonia gas?

  (b)   is it necessary to use an indicator in the preparation of sodium

tetraoxosulphate(VI) by neutralisation?

(c) is a solution of ammonium chloride in water acidic?

(d) are copper(II) trioxonitrate(V) and sodium tetraoxosulphate(VI) not prepared by double decomposition?

(e) is phenolphthalein not used as an indicator for the titration of a weak base with a strong acid?

(f) is it not advisable to dilute tetraoxosulphate(VI) acid by adding water to the acid?

(g) is the reaction of rubidium metal with water to give an alkali, more vigorous than that of sodium metal?

(h) does dry hydrogen chloride gas have no effect on dry blue litmus paper?

7. (a) Give two examples each of the following: hydrated, efflorescent, hygroscopic, deliquescent, acid and basic salts.

(b) List four methods of preparation of salts, and give an example of each method.

(c) Describe in detail how you would prepare samples of Na_2CO_3 and $PbSO_4$.

8. (a) Describe in detail how you would prepare a 0.1 mol dm^{-3} solution of Na_2CO_3, and titrate this with an approximately 0.1 mol dm^{-3} solution of HCl.

(b) About what volume of the acid will react with 50cm^3 of the base?

(c) What salts are present in solution when the following volumes of the acid have been added to 50cm^3 of the base: 50cm^3, 100cm^3, 150cm^3?

(d) Describe how you would obtain the normal salt formed between the acid and the base, from a dilute solution.