THE GOVERNMENT OF THE REPUBLIC OF THE UNION OF MYANMAR
MINISTRY OF EDUCATION

CHEMISTRY
GRADE 9

BASIC EDUCATION CURRICULUM, SYLLABUS AND TEXTBOOK COMMITTEE

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INTRODUCTION TO NEW EDITION

Chemistry may be generally defined as that branch of science which is concerned with the study of the composition, properties and structure of matter and the ways in which substances can change from one form to another or react with one another. Matter is anything which has mass and occupies space. It is present all over, and in the far corners, of the universe. The study of chemistry, therefore, implies the study of all processes which are going on all around us and even inside us.

The past decade has witnessed spectacular developments in chemistry and the related fields of scientific studies. Well over 3 million different substances or compounds are now known, a good number of them having been produced or synthesized by chemists in the laboratories. Thousands more are turned out synthetically every year in various parts of the world. All these new substances are being produced with the expectation that at least a number of them will benefit us and make our lives more comfortable. Most of the new substances used in medicine, agriculture and the textile and plastic industries, to mention a few, had their beginnings in the laboratory.

It would be an impossible task to learn the properties and reactions of each and every one of these substances. For a meaningful study of chemistry, the various component substances of the natural world are observed, measured and classified into appropriate categories. Scientific laws are then formulated to summarize the results of these numerous observations. Once the formulation of a scientific law is achieved, theories or models are devised to explain the law. Only those theories tested and proven by the entire scientific community survive to provide the basic foundations for the study of the science of chemistry. In the course of time a great many of the old established theories have been forced to the background by new developments resulting from a better understanding of the chemical processes and phenomena. Such new insights into the science of chemistry and the related scientific fields, coupled with the mountains of facts accumulated especially during the past decade or so, have had a profound effect on the study of chemistry.

Old ideas are revised or reformulated to blend in with the new. Still newer ideas are generated, developed and applied at a pace never before known in the history of chemistry.

These developments have prompted major changes in the way chemistry is studied and taught throughout the world. This is reflected not only in the new chemistry course contents but also in the course organizations and the methods of teaching and examining the subject. The modern day emphasis on practical work is one such prominent aspect of the new approach to the teaching of chemistry.

Fully aware of the new developments in the teaching of high school chemistry and taking cognizance of the fact that some of the students will need to have professional orientations on completion, the authors have undertaken the task of preparing two textbooks, one for the ninth standard and the other for the tenth standard of the Basic Education High School, in conformity with the revised high school chemistry syllabus. The course contents of these two textbooks taken together comprise the requirements for the Basic Education High School Chemistry Course.

The tenth standard (Grade-10) chemistry text is based on the knowledge the student has acquired from the study of the ninth standard(Grade- 9) chemistry text. Any student preparing for the Basic Education High School Examination should be familiar and proficient with the theoretical portions included in the ninth standard and the tenth standard chemistry texts as also with the experimental chemistry texts which accompany these books.

In this new edition, most of the chapters have been revised, summary and relevant questions for all chapters are added and updated. For the benefit of the students, four appendices are added and enhanced.: Appendix 1: Periodic table of the elements (updated 2003), Appendix 2: Table of the relative atomic numbers and atomic mass of the elements (including the new discovered web-elements), Appendix 3 Vapour pressure of water and Appendix 4 The Activity series.

The authors should like to acknowledge their gratitude to the Ministry of Education and to the Directorates of Basic Education and Higher Education for their help and encouragement during the writing of this book.

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CONTENTS

CHAPTER 1

SUBSTANCES AND THEIR CHANGES

1.1 Solids, Liquids and Gases

The difference between a solid and a liquid

A piece of iron is a solid. Water is a liquid. You know that a solid does not change in volume and shape, whatever container it is placed in.

We can therefore define a solid as follows:
A solid is a substance which has definite volume and definite shape.

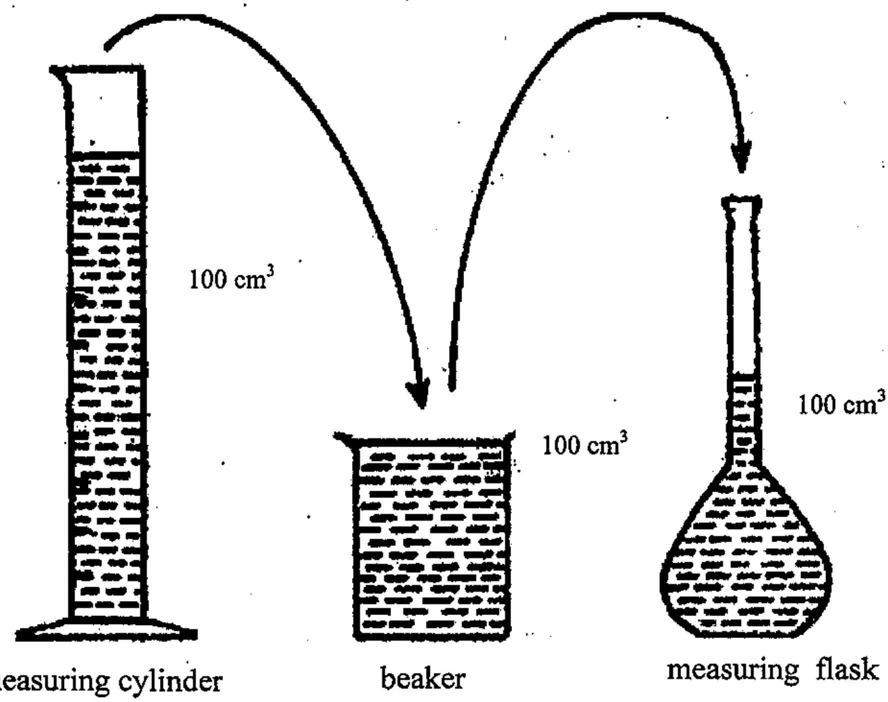


Fig. 1.1 Liquid and containers

$100\ cm^3$ of water in a measuring cylinder may be transferred successively into a $100\ cm^3$ beaker and then again into a $100\ cm^3$ measuring flask. (See Fig. 1.1) In each case there is no change in volume. But the shape of the liquid changes according to the shapes of the containers.

We can therefore define a liquid as follows:

> A liquid is a substance which has definite volume but no definite shape. It takes the shape of the container in which it is placed.

The difference between a gas and a solid or a liquid

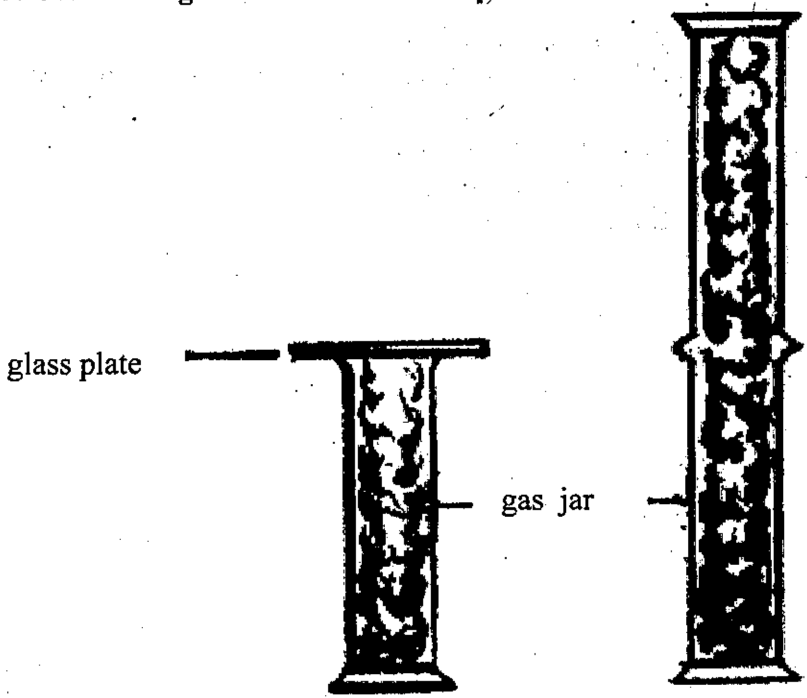


Fig. 1.2 Gas and containers

A few drops of bromine placed in a gas jar covered with a glass plate changes rapidly from a reddish brown liquid into a reddish brown vapour which fills the whole jar. If another jar of the same size is placed in an inverted position on the first jar as shown in Fig. 1.2 and the glass plate removed, both jars are soon filled with bromine vapour.

So we can define a gas as follows:

> A gas is a substance which has no definite volume and shape. It fills up every space it can enter and it takes up the volume and the shape of its container.

Three physical states

There are three states in which matter can exist - solid, liquid and gas. Every substance exists in one of the three states.

1.2. Change of States with Temperature

We can observe the changes which take place when ice is heated.

Some pieces of ice may be put in a beaker and its temperature measured with a thermometer. On heating the beaker the ice will slowly melt into water but the temperature will remain at 0°C so long as ice and water exist together.

After all pieces of ice have completely melted, the temperature will rise slowly to finally reach 100°C. At that temperature water will start to boil and change into steam, while the temperature remains steady at 100°C.

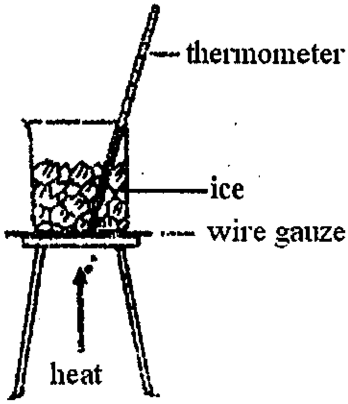


Fig. 1.3 Change of state of ice on heating

When steam is cooled, it changes into water at 100 °C. When water is cooled down to 0 °C it changes into ice.

Hence, water exists as solid (ice) at 0°C and below. It exists as liquid (water) between 0 °C and 100 °C. It exists as vapour (steam) at 100 °C and above.

In the same way other substances can exist in the solid, liquid or gaseous state depending on the temperature. We can summarize this observation diagrammatically as follows:

$$\boxed{\text{solid}} \underset{\text{cool}}{\overset{\text{heat}}{\rightleftharpoons}} \boxed{\text{liquid}} \underset{\text{cool}}{\overset{\text{heat}}{\rightleftharpoons}} \boxed{\text{gas}}$$

Melting and melting point

Melting is the change from the solid state to the liquid state, whereas freezing is the change from the liquid state to the solid state.

The temperature at which the solid and liquid of the same substance can exist together is called the melting point of the solid.

Evaporation and boiling point

The change from the liquid state to the vapour state is called vaporization whereas the reverse process is called condensation. Loss of liquid due to vaporization is called evaporation. When water is heated in an open vessel, its temperature slowly rises until 100 °C is reached when it begins to boil. Once the water boils, the temperature remains steady at 100 °C, even if more heat is applied, so long as the vessel is kept open.

The temperature at which the liquid boils in an open vessel is commonly called the boiling point of the liquid.

Sublimation

When solid iodine is heated in a test tube, it does not melt, but it changes directly into iodine vapour. When the vapour is cooled, it changes directly into solid iodine again. (See Fig. 1.4)

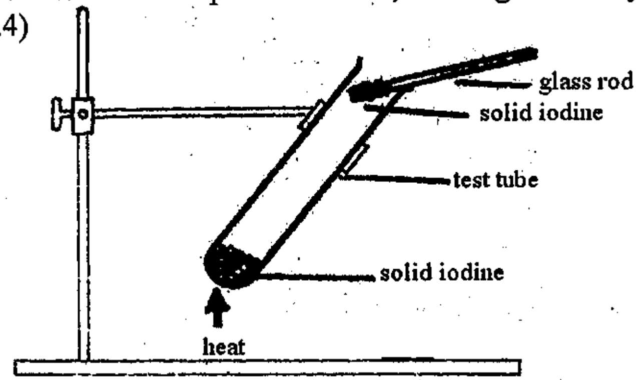


Fig. 1.4. Sublimation of iodine

Sublimation is a process in which a solid substance changes, without melting, directly into its vapour on heating.

Ammonium chloride and naphthalene can also sublime on heating.

However, these substances may exist in the liquid state under certain specific conditions.

1.3 Changes on Heating

Heating a nichrome wire

When a nichrome wire is heated, its temperature rises and it becomes red hot.

On cooling, it returns to its original form.

Thus on heating nichrome wire,

(a) the temperature and colour change, but

(b) no new substance is formed.

Burning a magnesium ribbon

When a magnesium ribbon is heated, its temperature increases, and then it burns brightly. After the burning, a white ash is left as residue.

Hence, on burning magnesium,

(a) heat and light are produced, and(b) a new substance (the ash) is formed.

1.4 Changes on Dissolving

Dissolving blue copper (II) sulphate crystals in water

Copper (II) sulphate dissolves in water and a blue solution is formed.

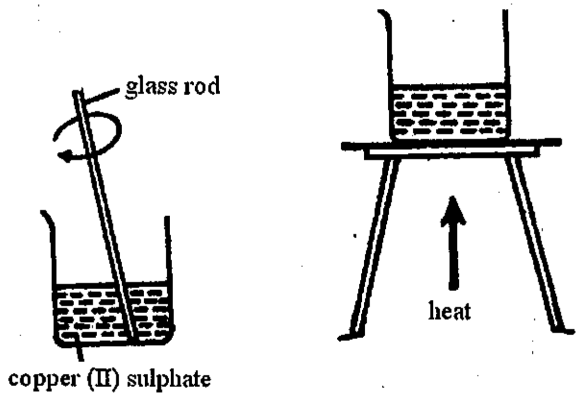


Fig. 1.5 Dissolution and recovery of copper (II) sulphate crystals

On boiling the solution, water evaporates slowly and finally blue crystals of copper (II) sulphate reappear(Fig. 1.5).

Hence, after dissolving copper (II) sulphate in water, the original substance may be recovered on evaporation.

Dissolving calcium in water
A small piece of calcium is placed in a dry test tube and the apparatus set up as shown in Fig. 1.6 (a).

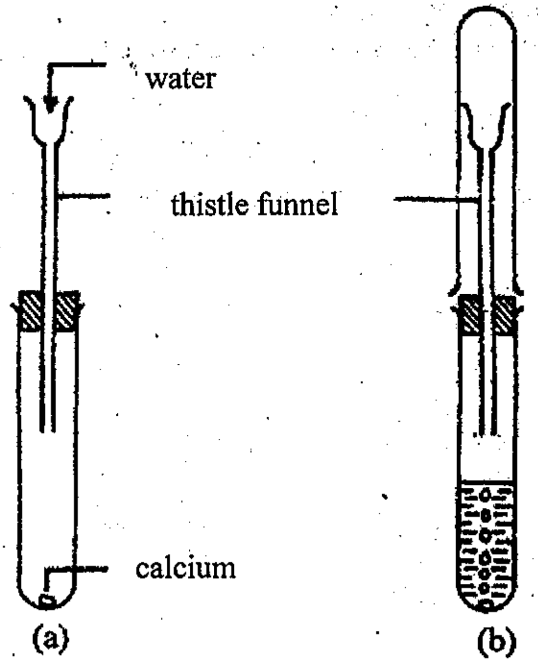


Fig. 1.6 Reaction of calcium with water

Water is poured down the thistle funnel, taking care to keep its level below the bottom of the funnel.

As soon as calcium makes contact with water, it gives off heat and hydrogen gas. The gas is collected in an inverted test tube as shown in Fig. 1.6 (b).

When the mouth of the test tube is placed over a lighted candle, the gas burns with a popping sound.

Test the aqueous liquid in the test tube with red litmus paper. The paper turns blue. Since pure water does not change the colour of litmus paper, the liquid in the test tube is different from water.

The following changes are observed on dissolving calcium in water,

(a) calcium disappears,
(b) heat and hydrogen gas are given off, and
(c) a new substance is formed.

1.5 Changes on Passing Electricity

Passing electricity through a copper rod

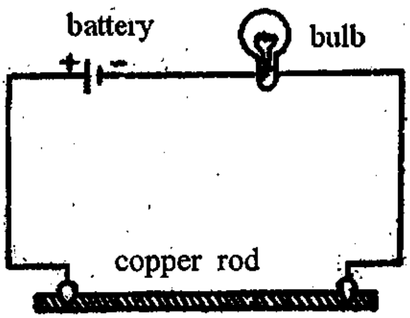


Fig. 1.7 Passing electricity through a copper rod.

When electricity is passed through a copper rod, included in a circuit as shown in Fig. 1.7, the electric bulb lights up. In this experiment

(a) the copper rod becomes electrified, but

(b) the copper rod retains its original properties when the passage of electricity is discontinued.

Therefore no new substance is formed.

Passing electricity into molten lead (II) iodide

Some lead(II) iodide is placed in a boiling tube and the apparatus set up as shown in Fig.1.8. The boiling tube is heated to melt the lead (II) iodide. Electricity is then passed through the molten lead (II) iodide.

Violet vapours of iodine are liberated from the positive carbon rod (anode). When the negative carbon rod (cathode) is examined after the experiment, metallic lead is seen.

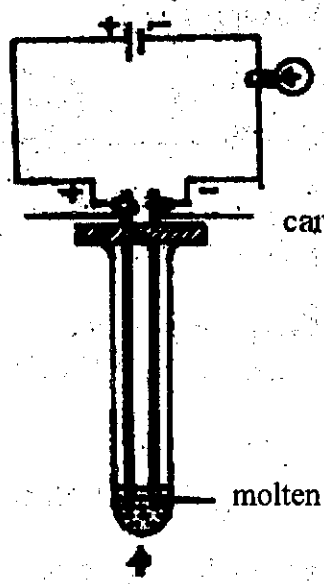


Fig. 1. 8 Passing electricity into molten lead(II) iodide

Thus on passing electricity through molten lead (II) iodide

(a) lead (II) iodide changes into lead and iodine and hence,

(b) new substances are formed.

1.6 Physical Change and Chemical Change

We have studied the changes involved on

(a) heating nichrome wire

(b) dissolving blue copper (II) shlphate crystals in water,

(c) passing electricity through a copper rod.

In these experiments, there may be a temporary change in colour, temperature, and state of the substances but no new substances are formed in the change.

A change of such nature is called a physical change.

In the experiments on

(a) burning magnesium ribbon,

(b) dissolving calcium in water,

(c) passing electricity into molten lead (II) iodide,

not only do the substances change in colour, temperature and state but they also change into a new substance or substances.

A change of such nature is called a chemical change.

To summarize, the most important observations regarding physical and chemical changes are the following:

A physical change is a change in which no new substances are formed.

A chemical change is a change in which one or more new substances are formed.

1.7 How to distinguish a Chemical Change from a Physical Change

To distinguish a chemical change from a physical change, we must know whether a new substance is formed in the change.

This can be ascertained as follows:

(a) By studying the physical properties of the substances; before and after the change:

(1) state (liquid, solid or gas)
(2) colour
(3) odour (smell)
(4) taste
(5) solubility in water (ability to dissolve in water)
(6) density
(7) melting point
(9) boiling point

Generally, the physical properties of different substances are different.

(b) By noting whether heat and light are produced: In physical changes the production of heat and light are usually not distinct; but in many chemical changes (such as burning of magnesium and reaction of calcium with water) heat and light are very distinctly produced.

(c) By examining the change in weight: In physical changes there is no change in weight because no new substances are formed. In a chemical change there is a change in weight of substance undergoing the change because it is converted into a new substance.

1.8 Changes and Energy

Every change, either physical or chemical, of a substance, is accompanied by a change of energy. As you know, heat, light, electricity, etc., are various forms of energy. In chemistry, we are more concerned with heat and chemical energy.

Chemical energy is a form of potential energy which exists in a substance. In chemical changes, this chemical energy is converted into heat and light, as you have seen in the burning of magnesium ribbon.

In some physical and chemical changes, energy is absorbed.

For example:

1. Heat is absorbed when ice melts into water.
2. Heat is absorbed when water changes into water vapour.
3. Heat is absorbed when iodine sublimes into iodine vapour.

4. Electrical energy is utilized when lead (II) iodide changes into lead and iodine.

In some physical and chemical changes, energy is evolved.

For example:

1. Heat is evolved when water vapour changes into water.
2. Heat is evolved when water changes into ice.
3. Heat and light are evolved when magnesium ribbon burns in air.
4. Heat is produced when calcium reacts with water.

1.9 Chemical Reaction

Chemical reaction

The process of undergoing a chemical change is called a chemical reaction.
(Give an example from the experiments you have studied.)

Reactants

A substance or substances that take part in a chemical reaction are called reactants.
(Name the reactants in the reaction of calcium and water.)

Products

A new substance or substances that are formed in a chemical reaction are called products.
(What are the products when electricity is passed through molten lead (II) iodide?)

Word equation for a chemical reaction

A chemical reaction can be described by a word equation. The names of the reactants are written on the left hand side, and the names of the products are written on the right hand side of the equation.

Reactant (s) = Product (s)

(a) When magnesium burns in air, it combines with oxygen of the air forming magnesium oxide. Here magnesium and oxygen are the reactants and magnesium oxide is the product.

magnesium + oxygen = magnesium oxide

(b) When calcium reacts with water, calcium hydroxide and hydrogen are formed. Write an equation in words for this reaction.

(c) Write an equation in words for the chemical change that takes place when electricity is passed through molten lead (II) iodide.

SUMMARY

In this chapter, the nature of substances and their changes is introduced first. Then the substances are classified into solids, liquids and gases and their differences are also explained. In addition, changes from one form to another have been illustrated. The factors that govern the physical and chemical changes are also

explained with examples. Questions and problems are given at the end of the chapter for exercise and practice.

Questions and problems

1. Describe an experiment to illustrate the difference between a solid and a liquid.
2. Describe an experiment to show the difference between a liquid and a gas.
3. The following scheme represents the general statement for the existence of three physical states.

$$\text{solid} \underset{\text{cool}}{\overset{\text{heat}}{\rightleftharpoons}} \text{liquid} \underset{\text{cool}}{\overset{\text{heat}}{\rightleftharpoons}} \text{gas}$$

   (a) Describe an experiment to show that the above statement is true.
   (b) Describe an experiment to show that the above statement has at least one exception.
4. Electricity can bring about both physical changes and chemical changes.
   (a) Describe an experiment in which electricity causes a physical change.
   (b) Describe an experiment in which electricity causes a chemical change.
5. Classify the following changes under two groups:
   (i) Physical changes (ii) Chemical changes
   (a) An egg being boiled until it becomes hard
   (b) Cooking of rice from rice grains
   (c) Sugar being dissolved in water
   (d) Green mangoes ripening
   (e) Freezing of water to ice
   (f) A solid on heating producing two gases
   (g) A solid on heating changing into a gas ; when the gas is cooled down its changing into a solid which has the same properties as the original solid before heating
   (h) Rotting of vegetables
6. In the following changes :
   (i) Is energy absorbed or evolved? (ii) Is a new substance formed?
   (a) Melting of ice
   (b) Burning of wood/paper
   (c) Heating of the filament in an electric bulb
   (d) Rusting of iron
7. Is dissolving table salt in water a physical change or a chemical change? Give reasons for your answer.
8. When calcium is dissolved in water, heat is produced and calcium hydroxide and hydrogen are formed. Is this change physical or chemical? What are the reasons for your answer?

9. Write a chemical equation in words for the change that takes place in question number 8.

10. When mercury (II) oxide is heated, mercury and oxygen are formed. Write a chemical equation for this change.

11. Write TRUE or FALSE for the following statements.
(a) Freezing of water to ice is a chemical change.
(b) Sugar being dissolved in water is a physical change.
(c) An egg being boiled until it becomes hard is a physical change.
(d) A solid on heating producing two gases is a chemical change.
(e) Rusting of iron is a physical change.
(f) Heating of nichrome wire produces a new substance.

12. Fill in the blanks with a suitable word or words.
(a) Heat is absorbed when .... melts into water.
(b) Heat is evolved when .... vapour changes into water.
(c) Heat is produced when .... reacts with water.
(d) Electrical energy is utilized when lead (II) iodide changes into .... and ....
(e) Heat is .... when water changes into ice.
(f) At 0°C water exists as .... and ....
(g) If the melting point of a substance is 20°C, its freezing point is ....
(h) .... is the the change from the vapour state to the liquid.

13. Select the correct word or words given in the brackets.
(a) Heat is (absorbed, evolved, produced) when ice melts into water.
(b) Heat is (absorbed, produced, evolved) when calcium reacts with water.
(c) Heat and light are evolved when magnesium ribbon burns in (nitrogen, air, carbon dioxide).
(d) Heat is absorbed when iodine sublimes into iodine (solid, liquid, vapour).
(e) Passing electricity through a copper rod is a (chemical, physical, static) change.
(f) Heat is (evolved, absorbed, produced) when iodine sublimes into iodine vapour.
(g) Water exists as (solid, liquid, gas) at 100°C and above.
(h) The change from the liquid state to the vapour state is called (evaporation, condensation, vaporization).

14. Match each of the items given in **list A** with the appropriate item given in **list B**.

List A	**List B**
(a) Melting of ice	(i) a physical change
(b) Burning of wood	(ii) a chemical change
(c) Sugar being dissolved in water	(iii) energy is evolved
(d) Rusting of iron	(iv) a new substance is formed
(e) Heating of the filament in an electric bulb	(v) energy is absorbed

15. Answer the following questions.

(a) What is the scientific term for cooling a vapour back into a liquid?

(b) What is the temperature at which a liquid water becomes a gas ?

(c) Give three substances which can sublime on heating.

(d) Describe one experiment on sublimation with diagram.

(e) Explain the terms ' reactant' and ' product'.

(f) Write the name of an experiment for chemical change .

(g) A liquid begins to freeze at -5°C. What is the melting point of that substance?

(h) When calcium is dissolved in water gas "X" is formed .

(i) Is this change physical or chemical ?

(ii) Name gas "X".

(iii) Is new substance formed ?

(iv) Write equation in words for this reaction.

(v) Name the reactants and products.

(i) Select the correct answer.

A chemical change takes place when

(i) milk sours.

(ii) a shirt is ironed.

(iii) electricity flows through a copper wire.

CHAPTER 2

ELEMENT, COMPOUND AND MIXTURE

In this chapter we shall classify elements into metals and non-metals. We shall also classify compounds into acids, bases and salts.We shall further try to differentiate between a compound and an element, and also between a compound and a mixture.

2.1 Substances that break down in chemical changes

Heating mercury (II) oxide

When mercury (II) oxide is heated in a test tube, it breaks down into mercury and oxygen.

mercury (II) oxide ⟶ mercury + oxygen

Mercury and oxygen are therefore, the components of mercury (II) oxide.

Passing electricity through lead (II) iodide

On passing electricity through molten lead (II) iodide, it breaks down into lead and iodine.

lead (II) iodide ⟶ lead + iodine

Hence, lead and iodine are the components of lead (II) iodide.

Some substances that are apparently single substances, such as mercury (II) oxide and lead (II) iodide, break down into other substances on undergoing chemical changes.

2.2 Elements : Substances that do not break down into other substances

If we heat individually, mercury, oxygen, lead or iodine in absence of air, we shall notice that these substances do not change into other substances.

Passage of electricity through lead or mercury also does not change them into other substances.

Hence, mercury, lead, oxygen and iodine cannot be broken down further into other substances by chemical means.

Substances like mercury, lead, oxygen and iodine, which cannot be broken down into simpler substances by chemical means, are called elements.

An element is a substance which cannot be broken down into simpler substances by chemical means.

Natural elements

There are 118 known elements which occur naturally, either in the free or combined state.

Some elements are solids such as copper, iron, zinc, silver, gold, carbon and phosphorus. Some elements are liquids. They are mercury and bromine. Some elements are gases such as oxygen, nitrogen, hydrogen and chlorine.

You can see some of them around you. Name ten elements which you can see in your school laboratory.

Some elements are man-made in the laboratory. Refer to the Periodic and Atomic Mass Tables for information regarding these elements.

2.3 Compounds

Combination of elements

When mercury is heated in air it combines with oxygen of the air to form mercury (II) oxide.

$$\text{mercury} + \text{oxygen} \xrightarrow{\text{heat}} \text{mercury (II) oxide}$$

When lead dust and iodine are heated together, they combine to form lead (II) iodide.

$$\text{lead} + \text{iodine} \xrightarrow{\text{heat}} \text{lead (II) iodide}$$

Hence, mercury (II) oxide and lead (II) iodide are composed of two elements each. Mercury (II) oxide and lead (II) iodide are called compounds.

Combination of elements by weight

When a known weight of mercury is heated in excess air contained in an enclosed vessel it will change to mercury (II) oxide. The mercury (II) oxide formed can be weighed. The difference in weight between the mercury and the mercury (II) oxide is the weight of oxygen.

This experiment may be repeated with a different weight of mercury and the weight of oxygen which has combined with this new weight of mercury can be found out.

Such experiments will show that mercury and oxygen always combine in a fixed proportion by weight to produce the new substance mercury (II) oxide.

Similarly you can find out that lead and iodine always combine in a fixed proportion by weight to form lead (II) iodide.

We can therefore, define a compound as follows:

> A compound is a substance formed by the combination of two or more elements in a fixed proportion by weight.

The difference between an element and a compound

The most obvious difference is that an element cannot be broken down into other substances by chemical means whereas a compound can be broken down into other substances by chemical means.

2.4 Classification of Elements

On the basis of their properties, elements may be classified into two groups, metals and non-metals.

(1) General properties of metals and non-metals

General properties of metals	*General properties of non-metal*
1. Metals show metallic lustre.	1. Non-metals do not show metallic lustre.
2. Metals have high density.	2. Non-metals have low density.
3. Most of the metals are malleable and ductile.	3. Non-metals are usually brittle.
4. Metals are good conductors of heat and electricity.	4. Most of the non-metals are poor conductors of heat and electricity.

Malleability: A substance is malleable if it can be beaten or pressed into new shapes. (Gold is highly malleable.)

Ductility: A substance is ductile if it can be drawn out in the form of a wire. (Gold is highly ductile.)

(2) Some important metals and their physical states

Some important metals and their physical states at ordinary temperature are listed below.

(1) Potassium	(solid)	(10) Iron	(solid)
(2) Sodium	(solid)	(11) Lead	(solid
(3) Calcium	(solid)	(12) Tin	(solid)
(4) Barium	(solid)	(13) Copper	(solid)
(5) Magnesium	(solid)	(14) Mercury	(liquid)
(6) Manganese	(solid)	(15) Silver	(solid)
(7) Aluminium	(solid)	(16) Gold	(solid)
(8) Zinc	(solid)	(17) Platinum	(solid)
(9) Chromium	(solid)		

All metals, except mercury, are solids at ordinary temperature. Mercury is a liquid at room temperature. Since it is a liquid, it is not malleable and ductile. But it has other metallic properties.

Some important non-metals and their physical states

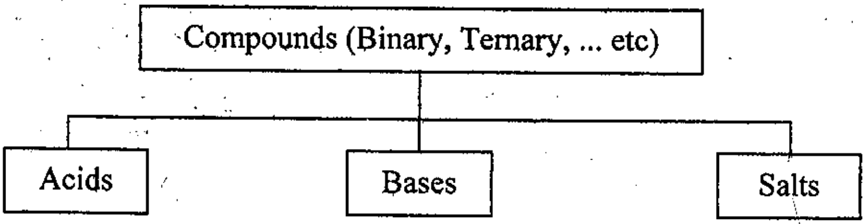
Some important non-metals and their physical states at ordinary temperature are listed below.

(1) Carbon	(solid)	(6) Nitrogen	(gas)
(2) Sulphur	(solid)	(7) Chlorine	(gas)
(3) Phosphorus	(solid)	(8) Bromine	(liquid)
(4) Oxygen	(gas)	(9) Iodine	(solid)
(5) Hydrogen	(gas)		

2.5 Classification of compounds

Two or more elements may combine together to form compounds. Some compounds occur naturally. Some compounds are man-made in laboratories.

The compounds can be classified in various ways. One useful and convenient method of classification is described as follows :



Binary compounds

Mercury (II) oxide is a compound of two elements, namely, mercury and oxygen.
Lead (II) iodide is a compound of two elements, namely, lead and iodine.

The compounds formed by the combination of two elements are called binary compounds.

A binary compound containing a metal and oxygen is called a metal oxide.
A binary compound containing a metal and chlorine is called a metal chloride.
A binary compound containing a metal and sulphur is called a metal sulphide and etc.
The common binary compounds which can be studied in the school laboratory are ;

oxide	chloride	bromide
calcium oxide silicon dioxide (silica or sand) copper (II) oxide magnesium oxide	sodium chloride (common salt) magnesium chloride silver chloride calcium chloride	lead (II) bromide potassium bromide
iodide	**sulphide**	**carbide**
lead (II) iodide potassium iodide	iron (II) sulphide copper (II) sulphide sodium sulphide	calcium carbide

Acids

The common acids which you can see in the school laboratory are :

Hydrochloric acid, Sulphuric acid, Nitric acid and Ethanoic acid.

Besides these common acids, you will comes across other acids such as phosphoric acid, carbonic acid, sulphurous acid and nitrous acid in this course.

All acids show some common properties. The characteristic properties of acids may be listed as follows:

(1) They have a sour taste.

(2) They turn blue litmus red.

Bases

Some metal oxides and all hydroxides are bases.

Some common bases which you can see in the school laboratory are :

Calcium oxide	(quicklime)
Potassium hydroxide	(caustic potash)
Sodium hydroxide	(caustic soda)
Calcium hydroxide	(slaked lime)
Ammonium hydroxide	

A common property of the bases is that they can turn red litmus blue. They also have a bitter taste.

Salts

Salts are formed when acids and bases neutralize one another. Some of the more common salts are :

Sodium chloride	(table salt)
Potassium nitrate	(nitre)
Magnesium sulphate	(epsom salt)
Sodium carbonate	(washing soda)

2.6 Mixtures

To some lead (II) bromide crystals on a piece of paper some mercury (II) oxide may be added with stirring.

No distinct changes will be observed in the process of mixing.

No chemical change occurs by mixing lead (II) bromide with mercury (II) oxide.

Similarly, you may verify that there is no chemical change on mixing sand and salt.

We have observed that no chemical change takes place on dissolving copper (II) sulphate in water. (See Chapter (1) section 1.4)

The whole mass obtained by mixing lead (II) bromide and mercury (II) oxide is called a mixture. In the mixture, though the substances are mixed together, they do not chemically combine with each other. No chemical change takes place on mixing.

The solution of blue copper(II) sulphate is also called a mixture because there is no chemical change on dissolving this substance in water, though blue copper (II) sulphate crystals are no longer present as solid. We can see that it is dispersed throughout the solution by its blue colour.

A mixture, therefore, is defined as follows :

> A mixture consists of two or more substances which may be present in any proportion by weight. The constituents of the mixture do not combine chemically.

2.7 Classification of Mixtures

Heterogeneous mixtures

In a mixture of copper (II) sulphate and sodium chloride, we can see the particles of these two substances distinctly with the naked eyes or with the help of a hand lens or microscope.

Mixtures whose constituents can be seen separately on visual or microscopic examination are called heterogeneous mixtures.

(a) A mixture of different solids such as copper (II) sulphate crystals and sodium chloride is a heterogeneous mixture(Fig. 2.1).

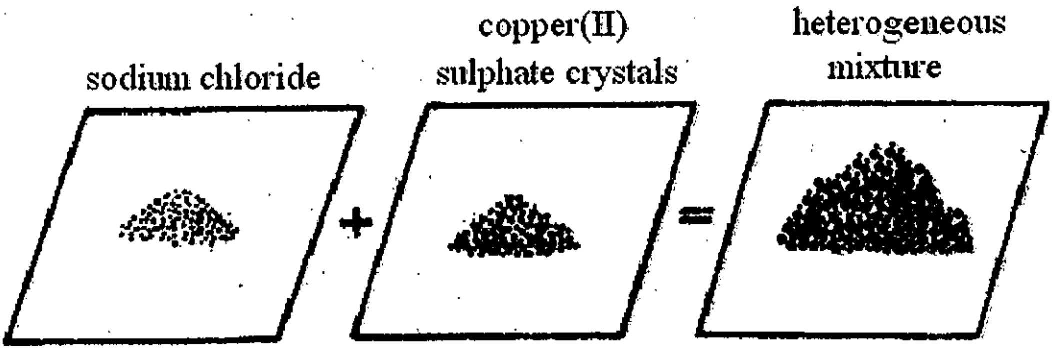
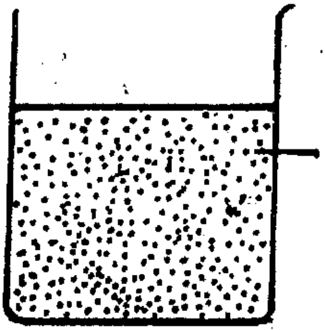


Fig. 2.1 Heterogeneous mixture of two solids

(b) A mixture of insoluble solids in a liquid such as chalk dust and water is also a heterogeneous mixture. On thorough mixing, a white liquid may be obtained which is actually a suspension of the chalk in water. On standing this mixture, the chalk particles will finally settle down in a separate layer(Fig. 2.2).



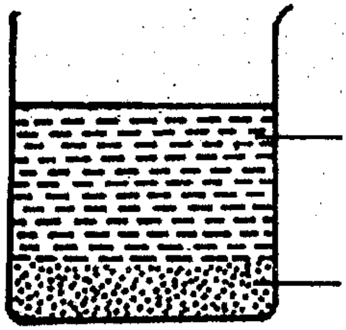


Fig. 2.2 Heterogeneous mixture of a solid and liquid

Other examples of heterogeneous mixtures may be observed in your home or laboratory. For example, milk is a heterogeneous mixture, it is actually an emulsion of oil in water. Other examples are the cosmetic creams such as face and hand creams and lotions.

Homogeneous mixtures

In a mixture, such as a solution of blue copper (II) sulphate crystals in water, we cannot see the constituents separately even under a high power microscope.

Mixtures whose constituents cannot be seen separately on visual or microscopic examination are called homogeneous mixtures.

Solutions obtained by dissolving

(1) solids in liquid ;
(2) liquids in liquid;
(3) gases in liquid ; and
(4) one gas mixed with another, are homogeneous mixtures.

2.8 Mixtures and Compounds

Some heterogeneous mixtures can be recognized even by the naked eyes. Homogeneous mixtures, however, cannot be recognized even on microscopic examination. Compounds are also homogeneous. So we cannot differentiate between a homogeneous mixture and a compound by visual examination. A mixture of iron filings and sulphur powder, and their compound, iron (II) sulphide, serve as a good example to illustrate the differences between a mixture and a compound.

The properties of iron filings and sulphur

Some of the properties of iron filings and powdered sulphur are shown in Table 2.1.

Table 2.1 Properties of iron filings and sulphur powder

	Properties of iron filings		*Properties of sulphur powder*
1.	Metallic grey colour.	1.	Yellow colour.
2.	No smell.	2.	Characteristic smell of sulphur.
3.	Attracted by a magnet.	3.	Not attracted by a magnet.
4.	Can react with hydrochloric acid to produce hydrogen gas which has no smell.	4	Does not react with hydrochloric acid.
5.	Does not dissolve in carbon disulphide.	5.	Dissolves in carbon disulphide.

Before we compare the essential differences between a mixture of iron and sulphur and iron (II) sulphide, we shall first consider how the mixture and the compound may be prepared.

Preparation of iron-sulphur mixture

Mix iron filings with sulphur powder, in any proportion by weight.

Examine the properties of iron sulphur mixture.

The properties of iron-sulphur mixture are shown in Table 2.2.

Preparation of iron-sulphur compound (iron (II) sulphide)

Mix iron filings with sulphur powder in the proportion of 7 parts by weight of iron to 4 parts by weight of sulphur. Heat the mixture in a test tube.

Chemical reaction takes place and the whole mixture glows red even when the tube is removed from the flame. This is the indication of a chemical change. After the reaction, a compound, called iron (II) sulphide, is formed. (Grind the iron (II) sulphide and examine it. Its properties are shown in Table 2.2)

Table 2.2 Properties of iron-sulphur mixture and iron (II) sulphide

Iron-sulphur mixture	*Iron-sulphur compound*
1. No chemical reaction takes place on mixing.	1. A chemical reaction takes place.
2. Separate particles of iron and sulphur can be seen.	2. A new substance is formed and no iron or sulphur particles can be seen in the substance.
3. The colours of iron and sulphur do not change.	3. The colour of the new substance is black.
4. A magnet can attract iron filings, leaving sulphur powder behind.	4. A magnet cannot attract the particles of the new substance.
5. On dissolving in carbon disulphide sulphur dissolves, leaving the iron filings behind.	5. The new substance does not dissolve in carbon disulphide.
6. The iron filings from the mixture react with hydrochloric acid to produce hydrogen.	6. The new substance reacts with hydrochloric acid and a gas with the rotten egg smell is given off.
7. Iron filings and sulphur powder may be mixed in any proportion by weight.	7. In the new substance iron and sulphur combine in a fixed proportion by weight.
8. The components can be separated by either a magnet or carbon disulphide.	8. No physical method can separate iron and sulphur from the new substance.

From the comparison of iron-sulphur mixture and iron-sulphur compound (iron (II) sulphide), some of the differences between mixtures and compounds can be described as follows :

Table 2.3 Differences between mixtures and compounds

Mixture	*Compound*
1. No chemical reaction takes place on mixing.	1. Chemical reaction takes place in the formation of a compound.
2. No new substance or substances are formed in a mixture.	2. A new substance or substances are formed.
3. Most of the properties of the constituents do not change in a mixture.	3. The properties of the compound are different from those of the substances from which it was formed.
4. The constituents of a mixture may be present in any proportion.	4. The constituents of a compound are present in a fixed proportion by weight.
5. A mixture can be separated into its constituent substances by mechanical or physical means such as the use of a magnet or dissolving out one constituent by a suitable solvent.	5. A compound cannot be broken down into its constituent elements by physical means.

How can you know that a substance is a mixture or a compound?

In considering whether a substance is a mixture or a compound, each of the following characteristics should be considered.

(1) Whether the substance is heterogeneous or homogeneous.
If the substance is heterogeneous, it is definitely a mixture. If it is homogeneous, further considerations are necessary.

(2) Whether the substance has properties which are the average of its constituents or whether the substance has different properties from its constituents.
If the substance has properties which are the average of those of its constituents, it may be a mixture. If the substance has properties entirely different from those of its constituents, it may be a compound.

(3) Whether the constituents of the substance are present in any proportion by weight, or whether they are present in a fixed proportion by weight.
If the constituents are present in any proportion by weight, the substance is a mixture. If the constituents are present in a fixed proportion by weight, the substance is a compound.

(4) Whether the constituents of the substance can be separated by mechanical or physical means or whether the use of chemical means is necessary. If the constituents of the substance can be separated by physical means, the substance is a mixture. If it can be separated only by chemical means, the substance is a compound.

SUMMARY

This chapter deals with elements, compounds and mixtures. The definitions of elements, compounds and mixtures are firstly presented . The classification of elements into metals and non-metals is explained. Compounds are classified into binary compounds, acids, bases, and salts are also given. In addition, the classification of mixtures into homogenous and heterogeneous types has been exemplified. Finally, the main differences between compounds and mixtures have been explained with appropriate examples such as iron filings and sulphur powder mixture and iron(II) sulphide compound.

Questions and Problems

1. When attempts are made to break down substance A by chemical methods the same original substance is always formed. Is substance A an element or a compound?
2. Barium is an element. Do you think new elements will be formed when barium is broken down by chemical means?
3. How can you determine that a given substance is an element?
4. When a substance is broken down by chemical means, two substances with different properties are formed. Is the original substance an element or a compound?
5. How can you determine that a given substance is a compound?
6. An element is light, solid and brittle. It cannot conduct heat and electricity easily. Decide whether the element is metallic or a non-metallic. Give reasons.
7. Give reasons why gaseous elements are non-metallic.
8. A substance contains three components. In four different samples of the substance, it was found that the three components are present in four different proportions by weight. Is the substance a mixture or a compound? Give reasons.
9. Pure water is a compound of two elements, namely hydrogen and oxygen. In pure water the proportion by weight of hydrogen and oxygen is 1:8. On distilling rain water, we get pure water. What is the proportion by weight of hydrogen and oxygen in the rain water after distillation?
10. There are two substances A and B. These two substances contain the same components x and y.
    (a) Three different samples of the substance A show the following compositions by weight.
    In sample 1, x = 5 g and y = 15 g
    In sample 2, x = 15 g and y = 45 g
    In sample 3, x = 6 g and y = 18 g

(b) Three different samples of the substance B show the following compositions by weight.

In sample 1, x = 2 g and y = 6 g

In sample 2, x = 3 g and y = 6 g

In sample 3, x = 4 g and y = 6 g

Determine which substance, A or B, is a compound and which substance, A or B, is a mixture.

11. Blue copper (II) sulphate crystals are dissolved in water to get a solution.
    (a) Does a chemical change take place on dissolving?
    (b) Does the colour of the copper (II) sulphate change on dissolving?
    (c) Say whether the solution is homogeneous or heterogeneous.
    (d) Can you change the proportion of blue copper (II) sulphate crystals and water in the solution?
    (e) By which means, physical or chemical, can you separate the blue copper (II) sulphate crystals from water?
12. Determine whether a solution of table salt in water is a mixture or a compound. Give reasons.
13. Write TRUE or FALSE for the following statements.
    (a) Substances that do not break down into other substances are called elements.
    (b) Some elements are not made by man in the laboratory.
    (c) An element is a substance which can be broken down into simpler substances by chemical means.
    (d) A compound is a substance formed by the combination of two or more elements in a fixed proportion by weight.
    (e) If the constituents are present in any proportion by weight the substance is a mixture.
    (f) A binary compound containing a metal and chlorine is called a metal chlorate.
    (g) Milk is a homogeneous mixture.
    (h) Different elements combine in any proportion by weight to form a compound.
14. Fill in the blanks with a suitable word or words.
    (a) Mercury combines with .... to form mercury (II) oxide.
    (b) Carbon, sulphur and phosphorous are ....
    (c) Salts are formed when acids and bases .... one another.
    (d) The solution of blue copper (II) sulphate is a mixture of copper (II) sulphate crystals .... in water.
    (e) All metals except mercury are .... at ordinary temperature.
    (f) Some metal oxides and all .... are bases.

(g) Mixtures whose constituents can be seen separately on visual or microscopic examination are called ....

(h) The compounds formed by the combination of two elements are called .... compounds.

15. Select the correct word or words given in the brackets.

(a) In a (an) (mixture, compound, element) a new substance or substances is (are) formed.

(b) Metals have (high, low, medium) melting points.

(c) A compound is a substance formed by the combination of two or more (substances, elements, solids) in a fixed proportion by weight.

(d) No (physical, distinct, chemical) change occurs by mixing lead (II) bromide with mercury (II) oxide.

(e) An element is a substance which cannot be broken down into simpler substances by (physical, chemical, biological) means.

(f) The common property of all acids has (bitter, sweet, sour) taste.

16. Match each of the items given in **list A** with the appropriate item given in **list B**.

List A	**List B**
(a) Acid	(i) can turn red litmus blue
(b) Non-metals	(ii) is a homogeneous mixture
(c) Mercury (II) oxide	(iii) are poor conductors
(d) Copper (II) sulphate crystal dissolved in water	(iv) can turn blue litmus red
(e) Bases	(v) is a compound of two elements

17. Answer the following questions.

(a) Name the metallic element in the liquid state at room temperature.

(b) Write down the characteristic properties of acids and name two common salts.

(c) Which of the following substances are pure substances and which are mixtures ?

(i) iron (ii) sea water (iii) soil (iv) gold
(v) oxygen gas (vi) coke (vii) mercury (viii) salt

(d) You are given iron filings and sulphur powder.

(i) Which substance has metallic grey colour and no smell ?
(ii) Name the substance which is dissolved in carbon disulphide.

(e) From the following list of elements, Aluminium, Zinc, Phosphorus, Chlorine, Lead, Sulphur, Mercury, Nitrogen, Potassium, Oxygen, which of the elements are non-metals?

(f) When three elements , A, B, and C combine to form a compound, the properties of this new compound ABC are (i) Like those of A and B (ii)

27 Like those of A,B and C (iii) Unlike either A or B or C (iv) Similar to both A and C.

(g) Write down the name of four binary compounds.

(h) Write down five differences between mixture and compound .

****

CHAPTER 3

SOLUTION

In this chapter we shall study the changes that occur when substances are dissolved in liquids.

3.1 Solution, Solute and Solvent

We can start with a simple experiment.

Some water is taken in a test tube. A small amount of copper (II) sulphate crystals are added to it with stirring. All solids will disappear, transforming into a clear blue liquid state in which the copper (II) sulphate is dispersed to give a homogeneous mixture. It is said that copper (II) sulphate dissolved in water and a solution of copper (II) sulphate in water is formed (Fig. 3.1).

Solvent + Solute ⟶ Solution (homogeneous mixture)

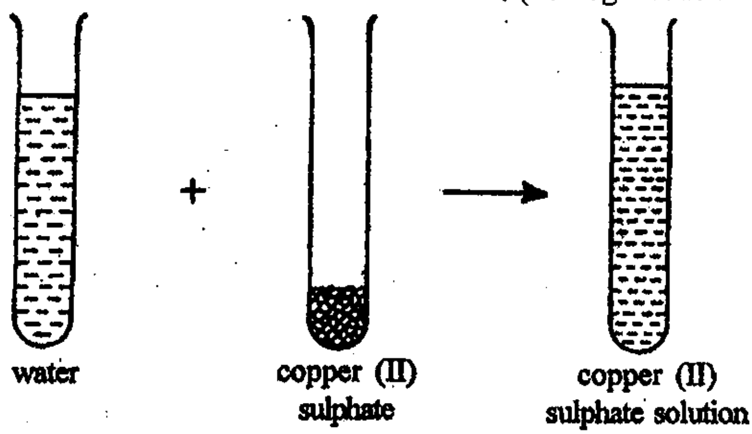


Fig. 3.1 Preparation of copper (II) sulphate solution

Solution
A solution is a clear homogeneous mixture obtained when a substance dissolves in a liquid.

Solvent
A solvent is a substance, mostly liquid, in which another substance dissolves to give a homogeneous mixture.

Solute
A solute is a substance which is dissolved in a solvent to give a solution.

In the above experiment, copper (II) sulphate is the solute. Water is the solvent. The homogeneous mixture obtained by dissolving copper (II) sulphate in water is the solution.

Distribution of solutes in solution

A solution is a homogeneous mixture of a solute and a solvent. In a solution the solute is uniformly distributed throughout the solution.

If 10 g of table salt is dissolved in 100 cm^3 of water and the solution thoroughly shaken, 1 g of table salt is distributed in every 10 cm^3 of the solution.

How many grams of table salt is present in 1 cm^3 of the solution?

3.2 Solubility and Insolubility of Substances

Some solids, such as copper (II) sulphate, sugar and common salt are soluble in water, but some solids such as sand, charcoal and chalk are insoluble in water. Some solids such as iodine are slightly soluble in water.

Two liquids that mix with one another in all proportions are termed miscible, while liquids that do not mix are termed immiscible.

Thus for example, alcohol mixes in all proportions with water; however, water and petrol are immiscible.

Generally, all gases are soluble in water. Some gases such as hydrogen chloride and ammonia are very soluble in water. Some gases such as oxygen, nitrogen and hydrogen are not very soluble in water.

Test for the solubility of a gas in water

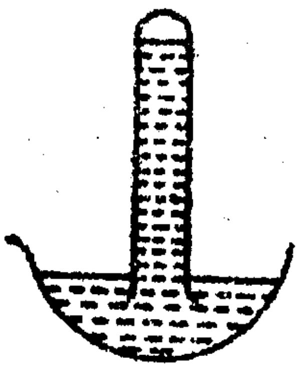


Fig. 3.2 Solubility of ammonia

A test tube is filled with ammonia gas and its mouth closed with the thumb. The closed end is then dipped into water contained in a porcelain basin. The thumb is then removed under the water and the test tube gently rocked. The water in the test tube will rise(Fig. 3.2). The height of water in the test tube indicates the solubility of the gas (How?). If the column of water in the test tube is high, the solubility of the gas also is high. If the water column is low the solubility of the gas is also low.

3.3 Saturation of the Solution

To prepare a saturated solution

Some solid is added to a certain volume of water in a beaker with stirring until all of it dissolved. Some more solid is then added with stirring. Addition of solid and stirring must be repeated until no more solid would dissolve at the experimental temperature. It must be made sure that the solution is truly saturated, by ensuring that no more solid dissolves in the solution no matter how long it is left or how vigorously it is stirred.

Effect of temperature on saturation

It is clear that a solution in which it is impossible to dissolve any more of the solute at a given temperature is said to be saturated with the solute at that temperature. If this saturated solution is heated, more solute can be dissolved in it. Therefore, the solution saturated at the lower temperature is no longer saturated at the higher temperature. (Note that there is no change in the amount of solvent.) Addition of some more solute will be required to make the solution saturated again at the higher temperature.

Therefore, with some exception, the solubility of the solute increases when the temperature of the solution is increased.

The solubility data of potassium nitrate at 60°C, 70°C and 80°C are given below.

The solubility of potassium nitrate

Temperature	°C	60	70	80
Solubility	g/100 g water	106	135	167

(There are exceptions. The solubilities of all gases and some solids such as calcium hydroxide and calcium sulphate decreases when temperature of the solution is increased.)

It is, therefore, necessary to specify the temperature at which the solution is saturated.

From the above discussion we arrive at the following important definitions regarding solutions.

Unsaturated solution: A solution in which more of the solute can dissolve at the given temperature is called an unsaturated solution.

Saturated solution: A saturated solution is one in which no more solute will dissolve at the given temperature, in the presence of excess solute.

We have learned that the amount of solute which dissolves in a given solvent depends on the amount of solvent used, the temperature of saturation, and the extent to which the substance dissolves. Hence to define solubility we must fix the amount of solvent used (say 100 grams) and the saturation of the solvent must be done at a fixed temperature.

Solubility of a substance at a given temperature is the mass in grams of the substance which will saturate 100 g of water, at that temperature.

Determination of the solubility of a solid in water at room temperature

The procedure involves the preparation of the saturated solution of the solid followed by the determination of the amount of solid dissolved in 100 g of water.

To prepare a saturated solution

About 100 cm^3 of distilled water is taken in a beaker. A small amount of solid is added to the water using a glass rod for stirring until the solid dissolves. Repeat the procedure of adding the solid and stirring until no more solid would dissolve at the experimental temperature.

To find the amount of solid dissolved in 100 g of water

A clean and dry evaporating basin is weighed. Using a clean and dry filter paper and a dry funnel, about half of the saturated solution is filtered directly into the weighed evaporating basin. The evaporating basin together with the saturated solution is again weighed. The solution is evaporated slowly. When most of the water has evaporated, the basin is placed on a steam bath (or a sand bath). The heating is continued until all the water has evaporated. The evaporating basin is then cooled in a desiccator and weighed together with the solid. The process of heating, cooling and weighing should be repeated until the weight becomes constant. From the results obtained the mass in grams of the solid which would have dissolved in 100 g of water is easily calculated. Refer to experimental chemistry for the determination of solubility of common salt.

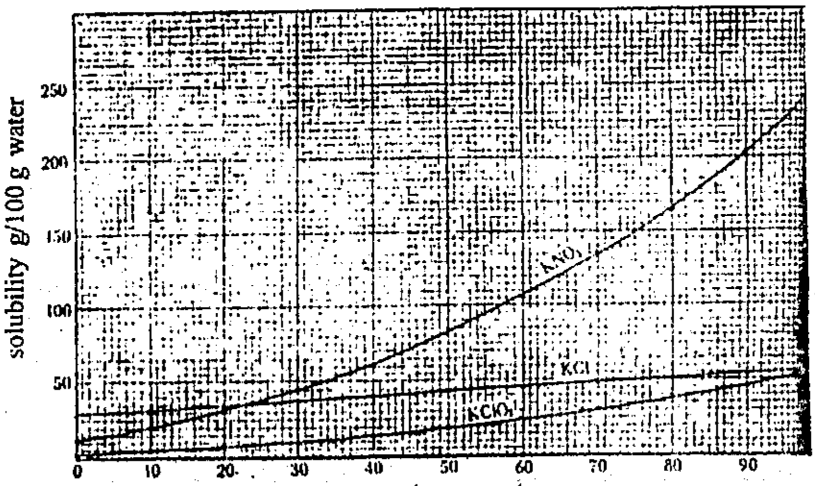
3.4 Solubility Graphs

Solubility data

Temperature °C	0	10	20	30	40	50	60	70	80	90	100
Solubility g/100 g water											
Potassium chloride	21.8	31.3	34.7	37.6	40.5	43.2	45.8	48.6	51.2	53.8	56.2
Potassium chlorate	3.3	5.1	7.3	10.1	14.0	18.5	24.0	30.2	37.4	46.0	56.2
Potassium nitrate	13.9	21.2	31.6	45.3	61.4	83.5	106	135	167	203	245

From the solubility data of the substances in the above table, a graph is drawn with temperature on the horizontal axis and solubility on the vertical axis as shown below.

Solubility graph



The use of solubility graphs

1. The curves indicate the changes in solubility of various salts with temperature.
2. Solubility of the substances between 0°C and 100°C can be found out from the graph.
   - (a) What is the solubility of potassium chloride at 40°C?
   - (b) What is the solubility of potassium chlorate at 70°C?
   - (c) What is the solubility of potassium nitrate at 60°C?
3. The solubility of the two substances at the same temperature can be compared and prediction can be made as to which substance will come out first when a solution saturated with these two substances is cooled down or is evaporated.

3.5 Supersaturated Solution

When a hot saturated solution is cooled to a given lower temperature, the solute in excess of the amount required to saturate the solution at the lower temperature usually comes out from the solution.

But this process does not always occur. Sometimes the excess solute can remain dissolved in the solution.

The solution that retains more solute than is required to saturate the solution at a given temperature is called a supersaturated solution.

To prepare supersaturated solution

For example, a supersaturated solution of hypo, $Na_2S_2O_3.5H_2O$ is readily made by directly heating the crystalline hypo when the sodium thiosulphate dissolves in its own water of crystallization to form supersaturated solution. On cooling this solution to room temperature, the excess hypo does not come out but remains in the solution. Because this solution contains more solute than it should at the room temperature, it is a supersaturated solution.

Supersaturation is an unstable condition, and if the solution is disturbed by adding a tiny crystal of the solute, all of the excess salt separates out and the remaining solution becomes an ordinary saturated solution.

3.6 Expressing the Amount of Solute in known Quantity of Solvent

In a saturated solution the fixed amount of solute is dissolved in a fixed amount of solvent at a given temperature. That is, in saturated solutions the ratio of the amount of solute to the amount of solvent does not vary. But in unsaturated solutions the ratio of the amount of solute to the amount of solvent may vary over a wide range.

The ratio of the amount of solute to the amount of solvent in solution can be expressed by the concentration of the solution.

A solution is referred to as being dilute or concentrated depending on whether it contains a small or large amount of solute in the same amount of solvent.

A Concentrated solution is a solution containing a comparatively large amount of the solute in the solution.

A dilute solution is a solution containing a comparatively small amount of the solute in the solution.

3.7 Crystals

When hot saturated solutions are cooled, the solid substances come out from the solution. These solids have glistening appearances and are bounded by regular flat faces forming characteristic shapes. These solids are called crystals. Most of the solids are crystals. Solids which have no natural characteristic shapes are not called crystals. The substances which have no natural crystalline shapes are called amorphous substances.

Shapes of crystals

The crystals of the same substance have the same crystalline form.
The crystals of different substance have different crystalline forms(Fig. 3.3).

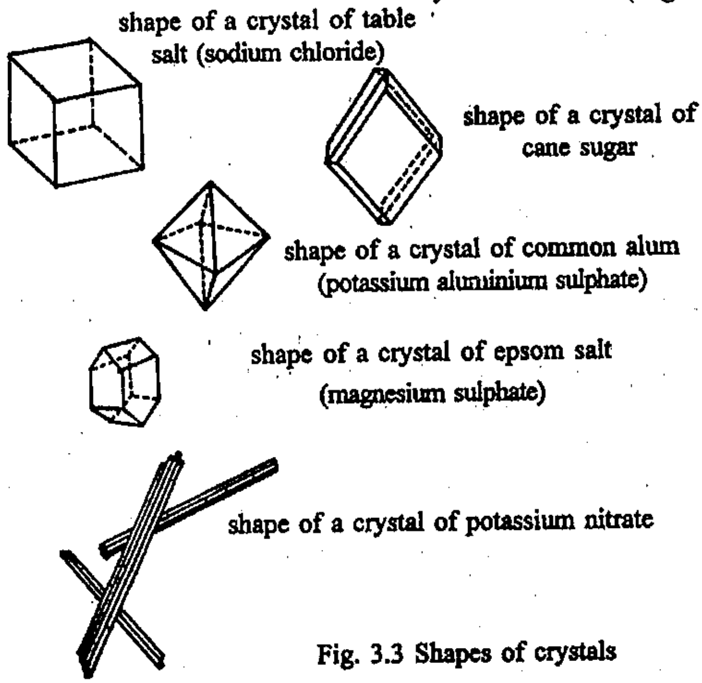


Fig. 3.3 Shapes of crystals

Action of heat on crystals

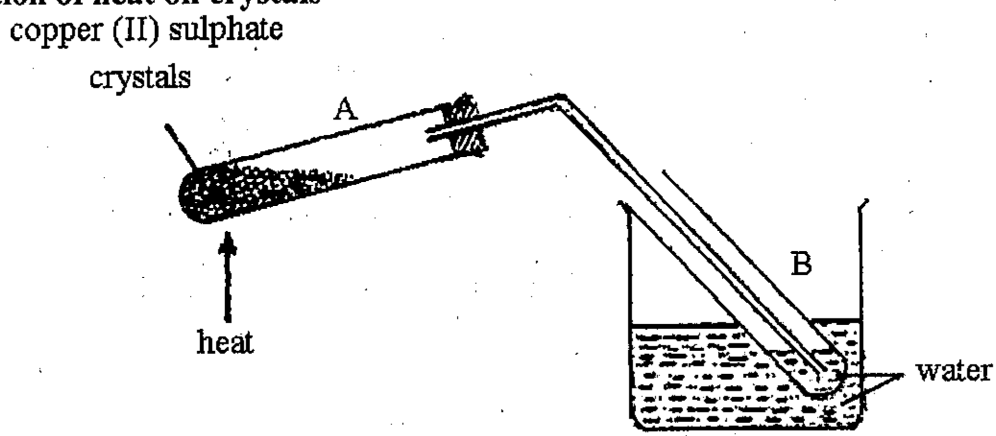


Fig 3.4 Heating copper (II) sulphate crystals

Some copper (II) sulphate crystals are placed in test tube A. The apparatus is set up as shown in Fig. 3.4 by placing tube B in water for cooling. The test tube A is heated. During heating copper (II) sulphate crystals gradually lose their blue colour and crystalline shape. Finally they all turn into a white powdery mass. Some colourless liquid is collected in test tube B. This liquid is water.

From this experiment it is clear that when copper (II) sulphate crystals are heated, the water from the crystals is lost and the resultant white mass in test tube A (copper (II) sulphate without water or anhydrous copper (II) sulphate) is obtained. When a small amount of water is added to the white anhydrous copper (II) sulphate, blue copper (II) sulphate reappears. So water is one of the components responsible for the colour and crystalline form of copper (II) sulphate crystals.

Water of crystallization: The water essential for the existence of the substance in the crystalline state is called water of crystallization.

Hydrate : Crystals containing water of crystallization are called hydrates.

Crystals without water of crystallization

On heating some table salt you will find that no water is given off. So, in the crystals of table salt there is no water of crystallization.

The names of some hydrated and some anhydrous crystalline substances are given below.

Crystalline substances with water of crystallization	**Crystalline substances without water of crystallization**
Copper (II) sulphate	Table salt
Magnesium sulphate	Potassium sulphate
Zinc sulphate	Iodine
Washing soda	

3.8 Making Crystals (Crystallization)

The crystals can be made in the laboratory by the following methods.

1. By cooling the saturated solution
2. By slow evaporation of the saturated solution.
3. By adding seed crystals to a saturated solution

By cooling saturated solutions

1. A saturated solution of the substance is prepared at high temperature.
2. The solution is allowed to cool.

(a) If the hot solution is rapidly cooled by immersing the container of the solution in water, the crystals are formed rapidly. But the size of the crystals will be small.

(b) If the hot solution is allowed to cool slowly by leaving the solution at room temperature for some time, the crystals are formed slowly. The crystals will be large in size.

By slow evaporation of saturated solutions

1. A saturated solution of the substance is prepared.
2. The solution is allowed to evaporate slowly at room temperature by leaving it in a container covered with a sheet of paper to prevent the entering of dust.

The water should evaporate slowly at room temperature and crystals of large size will be formed.

By using seed crystals

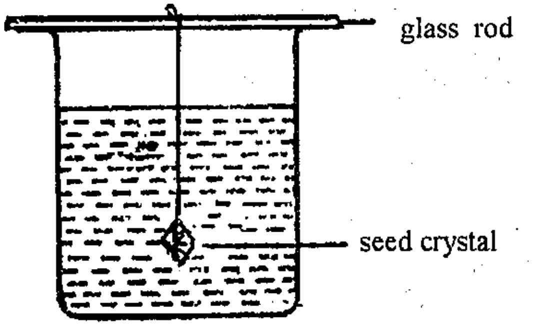


Fig. 3.5 Growing crystals

1. A saturated solution of copper (II) sulphate is prepared at room temperature. Filter the solution.
2. A perfect crystal of copper (II) sulphate is selected.
3. The selected crystal is suspended on a cotton thread in the saturated solution of copper (II) sulphate and left for some time (Fig. 3.5).
4. Solute will deposit on the seed crystal which will gradually grow to form a big crystal.

   This method is also used to prepare large crystals of common alum in the laboratory.

3.9 Efflorescence, Deliquescence and Hygroscopicity

Efflorescence

If washing soda (sodium carbonate) crystals are left in the air they lose some water of crystallization and the crystals transform into white powder which coats the surfaces of the crystals.

The process of losing all or part of the water of crystallization into the air from the hydrates is called efflorescence.

Deliquescence

When some substances are exposed to the air they absorb water from the air and become damp. This process is called deliquescence. If these substances are left exposed to the air for a long time deliquescense will finally cause all the solid to dissolve in the absorbed water. Examples of some common deliquescent substances are sodium hydroxide, calcium chloride, calcium nitrate and iron (III) chloride.

Hygroscopic substances

Some substances on exposure to the air absorb water from the air but may not dissolve in it. These substances are called hygroscopic substances. Examples of hygroscopic substances are copper (II) oxide, sodium nitrate and calcium oxide.

Drying agents

As deliquescent and hygroscopic substances can absorb water vapour from the air these substances are used to dry other substances.

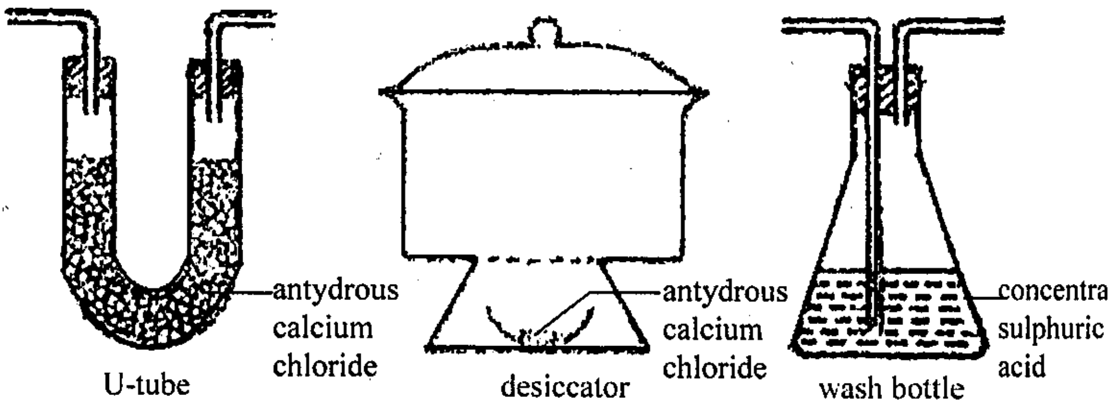


Fig. 3.6 Apparatus for drying

Anhydrous calcium chloride in a small bottle is kept inside a balance case to keep the air in it dry. Anhydrous calcium chloride in a U-tube is used to dry gases.

Calcium chloride is also used in desiccators to dry substances. Concentrated sulphuric acid in a wash bottle is used to dry gases which have no reaction with it. When used as contact drying agents these substances must have no reaction with the substance to be dried (Fig. 3.6).

SUMMARY

In this chapter, the concepts of solution, solvent and solute are clearly explained. The terms such as unsaturated solution, saturated solution, supersaturated solution and solubility have been elaborated. In addition, crystals of different substances are depicted and methods of making crystals are presented. The definitions of efflorescence, deliquescence, and hygroscopic terms are given.

Questions and problems

1. 20 g of a soluble substance is dissolved in water to form 100 cm^3 of the solution. 25 cm^3 of the solution is taken and evaporated to dryness. How many grams of the solid will be obtained?

2. A patient has to take medicine powder, 0.5 g at a time and four times a day. The medicine powder is water soluble. You are provided with 2.0 g of medicine powder, 40 cm^3 of water, and a bottle which can contain 40 cm^3 of liquid and has four marks that divides four equal parts.
   (a) How would you prepare the medicine?
   (b) How will you instruct the patient to take his medicine?

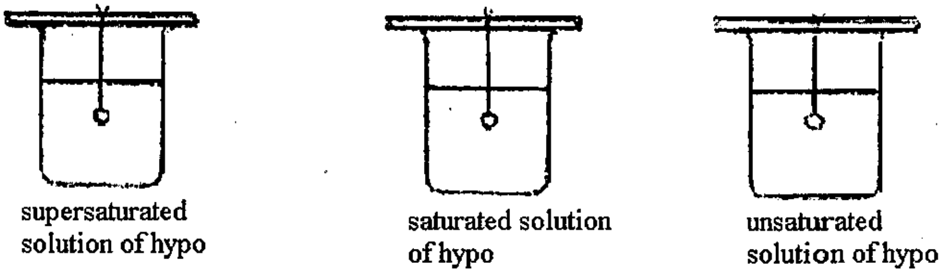
3. Three solid substances A, B and C are given to you. The substance C is practically insoluble in water. Excess amounts of the three substances are separately shaken with distilled water in a test tube. An equal amount of each solution is taken and evaporated to dryness. After evaporation a large amount of A is obtained and very little amount of B is obtained.
   (a) Which substance is very soluble?
   (b) Which substance is slightly soluble?
   (c) What do you expect to see after completely evaporating the liquid in which substance C is shaken?

4. Choose the proper words to complete the following statements.
   (a) When a solvent has dissolved the maximum amount of solute at a given temperature the solution obtained is said to be (saturated, unsaturated, supersaturated).

(b) The maximum amount of a solute that can be dissolved in a solvent at a given temperature is the amount which makes the solution (saturated, unsaturated, supersaturated) at that temperature.

5. Sugar is added into a glass of lime juice and stirred well until some sugar is left undissolved. The lime juice free from undissolved sugar is poured into a second glass. Do you think that the sweetness of the lime juice will increase if more sugar is added to the lime juice in the second glass? Give reasons.

6.



Three crystals of hypo are suspended in three solutions of hypo as shown in the diagram and are left to stand for some time.

(a) What change in the size of the crystal will occur in the supersaturated solution?

(b) What change in size of the crystal will occur in the saturated solution?

(c) What change in size of the crystal will occur in the unsaturated solution?

7. Three solutions of alum were separately placed in three beakers, A, B, and C. Three crystals of alum were separately suspended by means of thread in the solution in each beaker. As the time went by it was found that the crystal in beaker A, became smaller and smaller the crystal in beaker B, became larger and larger and the crystal in beaker C, did not change in size.

(a) Which solution is a saturated one?

(b) Which solution is a supersaturated one?

(c) Which solution is an unsaturated one?

8. There are two ways by which a saturated solution can be obtained from an unsaturated one at a given temperature. How will you do it both ways? There are two ways by which an unsaturated solution can be obtained from a saturated one at room temperature. How will you do it?

9. The solubility of solid A at 60°C is 24 and at 40°C is 14.

(a) What is the amount of the solid required to saturate 50 g of water at 60°C?

(b) What is the maximum amount of the solid that can be dissolved in 25 g of water at 60°C?

(c) How much solid will be obtained if 62 g of saturated solution at 60°C is evaporated to dryness?

(d) What will be the amount of saturated solution obtained at 40°C after dissolving 7 g of the solid in required amount of water?

(e) You have a saturated solution of 114 g at 40°C. How much more solid must be dissolved in this solution to make a saturated solution at 60°C?

(f) A saturated solution of the solid in 100 g of water at 60°C is cooled to 40°C. How much solid should come out of the solution?

(g) How much solid should come out of the solution if 124 g of the saturated solution of the solid at 60°C is cooled to 40°C?

10. The solubility of copper (II) sulphate at 60°C is 40 and 90°C is 67.5.

A saturated solution of copper (II) sulphate in 100 g of water at 90°C is cooled to 60°C. Calculate the amount of copper (II) sulphate which would come out of the solution.

11. The solubility data of sodium nitrate at various temperatures are given below.

Temperature °C	0	10	20	30	40	50	60	70	80	90	100
Solubility in g /100 g water	73	80	83	-	104	114	124	-	148	161	178

(a) Plot a solubility graph.

(b) What are the solubilities of sodium nitrate at 30°C and 70°C?

(c) You have a saturated solution of the substance in 100 g water at 50°C. How much more sodium nitrate should be dissolved to make a saturated solution at 70°C?

(d) You have a saturated solution of the substance in 100 g water at 90°C. How much solid should come out of the solution if the solution is cooled to 40°C?

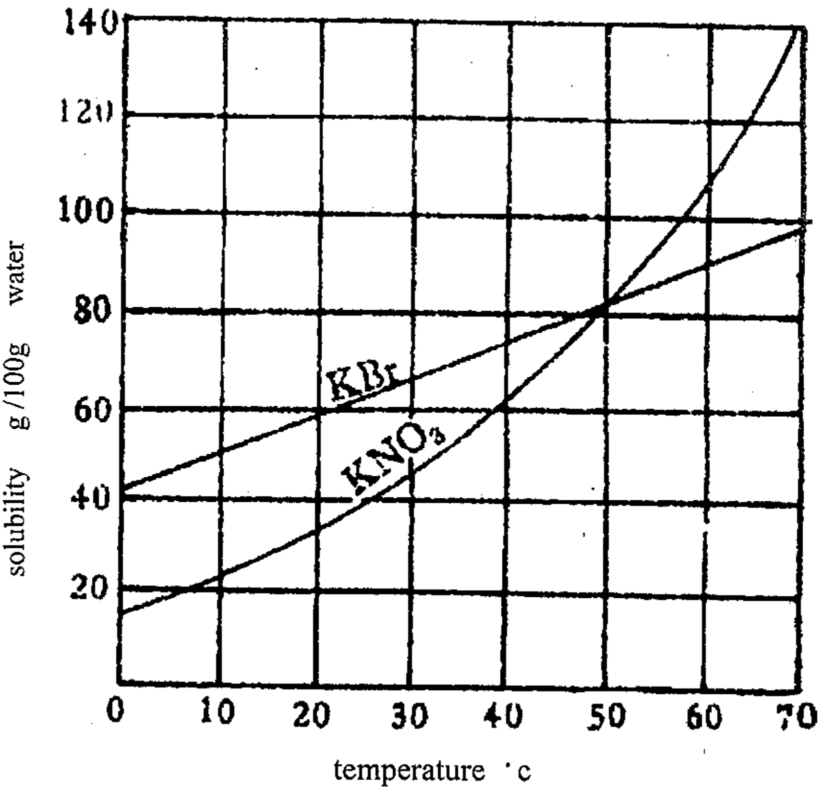
12. The solubilities (g/100 g water) of three substances at different temperatures are given below.

Temperature °C	0	10	20	30	40	50	60	70	80
Potassium chlorate	3.3	5.0	7.3	10.0	14.0	18.5	24.0	30.2	37.5
Calcium hydroxide	0.13	0.125	0.12	0.11	0.102	0.0	0.08	0.06	-
Sodium sulphate	5.0	9.0	20.5	41.0	48.0	47.0	45.0	44.0	43.0

(a) Plot the solubility curve of each substance.

(b) Describe the change in solubility with the temperature for each substance.

13.



Which substance is more soluble above 50°C and which substance is more soluble below 50°C?

14. Write TRUE or FALSE for the following statements.
   (a) Crystals containing water of crystallization are called hydrates.
   (b) A solution is a clear heterogeneous mixture obtained when a substance dissolves in a liquid.
   (c) A solute is a substance which is dissolved in a solvent to give a solution.
   (d) A saturated solution is one in which more solute will dissolve at the given temperature, in the presence of excess solute.
   (e) A solution in which more of the solute can dissolve at the given temperature is called an unsaturated solution.
   (f) Table salt is not a hydrate.
   (g) A solution is a homogeneous mixture of a solute and solvent.

15. Fill in the blanks with a suitable word or words.
   (a) Solubility of a substance at a given temperature is the mass in grams of the substance which will .... 100 g of water, at that temperature.
   (b) The solution that retain more solute than is required to saturate the solution at the given temperature is called a .... solution.
   (c) A .... solution is one in which no more solute will dissolve at the given temperature, in the presence of excess solute.
   (d) A .... is a substance, mostly liquid, in which another substance dissolve to give a homogeneous mixture.
   (e) A .... is a clear homogeneous mixture obtained when a substance dissolves in a liquid.
   (f) If the solvent is water, the solution formed is called an .... solution.
   (g) Calcium oxide is an example of .... substance.
   (h) Substances which dissolve in a particular solvent are said to be ....

16. Select the correct word or words given in the brackets.
   (a) A solution is a clear (homogeneous, heterogeneous, uniform) mixture obtained when a substance dissolves in a liquid.
   (b) A solution in which more of the solute can dissolve at the given temperature is called a (unsaturated, saturated, supersaturated) solution.
   (c) The water essential for the existence of the substance in the (solid, crystalline, amorphous) state is called water of crystallization.

(d) The process of losing all or part of the water of crystallization into the air from the hydrates is called (deliquescence, efflorescence, hygroscopicity).

(e) (Sodium carbonate, Washing soda, Sodium hydroxide) loses water of crystallization into the air.

(f) A (saturated, unsaturated, supersaturated) solution is a solution containing large amount of the solute in the solution.

(g) Hygroscopic substance is (copper (II) oxide, calcium oxide, sodium hydroxide).

17. Match each of the items given in list A with the appropriate item given in list B.

List A		List B	
(a)	The process of losing all or parts of the into the air from the hydrate is caled	(i)	The solution that retains more solute than is required to saturate the solution at a given temperature
(b)	Supersaturated solution	(ii)	Efflorescence
(c)	Hygroscopic substances	(iii)	Substances which have no natural crystalline shapes
(d)	Amorphous substances	(iv)	Apparatus for drying
(e)	Desiccator	(v)	Substances on exposure to air absorb water from the air but may not dissolve in it.

18. Answer the following questions.

(a) Hygroscopic substances are used as drying agents. Why?

(b) Give the name of two hydrates and two drying agents.

(c) Select the correct answer.

When sugar is dissolved in water,

(i) the solvent is sugar and the solute is water.

(ii) a chemical change takes place.

(iii) the boiling point of water decreases.

(iv) the resulting mixture is called a solution.

****

CHAPTER 4

THE LAWS OF CHEMICAL COMBINATIONS AND THE IDEA OF ATOMS

In Chapter 2 we have learned that elements in a compound combine in a definite proportion by weight. We can experimentally determine the proportions by which the elements combine to form a compound. One such experiment will be described using magnesium and oxygen from the air.

4.1. To find the Proportion by Weight of Magnesium and Oxygen in Magnesium Oxide

When a known weight of magnesium is heated in air, it combines with oxygen from the air to form magnesium oxide. The difference between the weights of magnesium and the oxide formed is the weight of oxygen. From the results, the proportion by weight of magnesium and oxygen in magnesium oxide can be calculated.

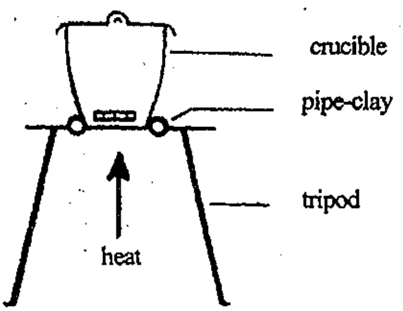


Fig. 4.1 Ignition of magnesium

A clean, dry crucible and its lid are first weighed. A length of magnesium ribbon (about 20 cm) is cleaned by rubbing it two or three times with sand paper. This magnesium ribbon is made into a tight coil and placed in the crucible which is then covered it with its lid.

The weight of the magnesium and crucible with its lid is then determined. The covered crucible is then placed on a tripod, using a pipe-clay triangle to support it. It is then carefully heated until the magnesium ignites (Fig. 4.1).

During heating, the lid must be lifted occasionally with a pair of tongs to let air in. Care must be taken not to allow the white fumes (magnesium oxide) to escape.

After the magnesium has completely burnt, the flame is removed and the crucible is cooled down. Then a drop or two of water is added and the crucible again heated. From this stage onwards, the crucible must be strongly heated to make sure that the combustion of the magnesium is completed.

Heating is then discontinued and the crucible cooled down. The crucible, lid and magnesium oxide are then weighed. The process of heating, cooling and weighing must be repeated until there is no more increase in weight.

Treatment of results

Weight of crucible and lid	=	W_1 g
Weight of crucible, lid and magnesium	=	W_2 g
Weight of crucible, lid and magnesium oxide	=	W_3 g
∴ Weight of magnesium	=	$(W_2 - W_1)$ g
∴ Weight of oxygen combined with magnesium	=	$(W_3 - W_2)$ g

Therefore,

Weight of magnesium	:	Weight of oxygen
$(W_2 - W_1)$	:	$(W_3 - W_2)$

The ratio of the two elements $(W_2 - W_1) : (W_3 - W_2)$ is consistently 3:2 .

Constancy of the composition by weight of magnesium oxide

At least three experiments should be done by using three different weights of magnesium.

The results of all these experiments will show that the ratio of the weight of magnesium to that of oxygen is always 3 : 2.

Therefore, irrespective of the starting weight of magnesium, the weight ratio of magnesium to oxygen in magnesium oxide is always constant.

Hence, it can be concluded that in magnesium oxide the magnesium and oxygen are always combined in a definite proportion by weight.

4.2 Comparison of the Total Weight of the Products of a Chemical Reaction with the Total Weight of the Reactants

We have learned that in chemical reactions the reactants change chemically. Now we may ask a question; " Is there any change in weight when substances react to form new products?"

The answer to this question may be obtained by comparing the total weight of the products of a reaction with the total weight of the reactants.

Some pairs of reactants are listed below. Use solutions of any pair of reactants for each experiment.

Magnesium sulphate and barium chloride
Silver nitrate and sodium chloride
Lead (II) nitrate and potassium iodide

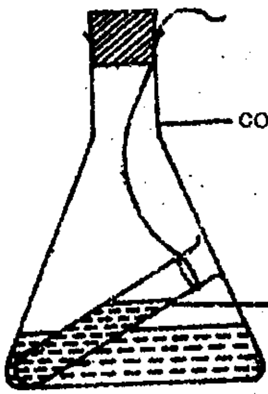


Fig. 4.2. Verification of law of conservation of mass

A conical flask is taken and one of the solutions of any pair given in the list is placed in it just to cover the bottom. A piece of thread is tied around the mouth of a small test tube which is half filled with the other solution of the pair. The test tube is held in position in the conical flask by means of the thread and tight fitting cork as shown in the diagram (Fig. 4.2).

The whole apparatus is weighed. Let the weight be W_1 g.

The flask is tilted so that the two solutions mix. A visible chemical change will take place. The cork ensures that all the substances remain inside the flask.

The whole apparatus is again weighed after the reaction. Let the new weight be W_2 g.

We can verify that W_1 g $=$ W_2 g

We can repeat the experiment with the solutions of the remaining pairs of reactants. In each of these reactions it is found that the weight of the reactants is equal to the weight of the products.

4.3. Constancy of Mass in Chemical Reactions

We have seen in the above discussions that in chemical reactions, the weight of the reactants is always equal to the weight of the products. So we can say that, in a chemical reaction, the total weight of the substances always remains the same. That is, the total mass remains constant throughout a chemical reaction.

This discovery leads to the idea that though substances change into other substances, the amount of matter in the substances, that is the mass, does not change. This is because no new matter is created, nor is any of the old matter destroyed, during the reaction. This is a law of nature known as the law of conservation of mass.

The law of conservation of mass states that:

> Matter can neither be created nor destroyed.

Hence the total mass of the substances before and after a reaction is always the same.

4.4 Constancy of Composition by Weight of the Compounds

Irrespective of the method used for its preparation, the components in a given compound always combine in definite proportions by weight. For Example, the weight proportion of hydrogen and oxygen in water is always 1:8. The proportion by weight of carbon and oxygen in carbon dioxide is always 3:8.

The law of definite proportions
The constancy of composition of compounds is another law of nature known as the law of definite proportions.

The law states :

> The same compound always contains the same elements in the same proportion by weight.

The importance of the above two laws
(1) These two laws, especially the law of definite proportion reinforce the idea that all substances are made up of very small particles or atoms.
(2) The law of conservation of mass is important because we can make use of it in writing chemical equations.

4.5 The Laws of Chemical Combinations and the Idea of Atoms

Why do elements combine in definite proportions by weight? Why does the total mass of the substances in a chemical reaction remain constant?.

These questions can be easily answered by assuming that :

> Matter is composed of extremely small particles called atoms.

We may further assume that :

> Atoms can neither be created nor destroyed in chemical reactions and that they combine in simple proportions to form small particles of compounds called molecules.

With these assumptions we can explain chemical reactions as described in the example below..

Combination of magnesium and sulphur

Magnesium is a metal. Sulphur is a non-metal. They combine to form magnesium sulphide on heating together.

magnesium + sulphur $\longrightarrow$ magnesium sulphide

It is found that the weight ratio of magnesium to sulphur in magnesium sulphide is always 3:4.

To explain this we will assume that ;

(a) Magnesium is composed of magnesium atoms. These atoms have the same mass.

(b) Sulphur is composed of sulphur atoms. These sulphur atom also have the same mass among them.

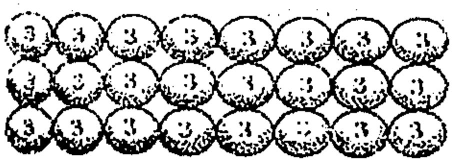
(c) The mass ratio of one magnesium atom to one sulphur atom is 3:4.

(d) One atom of magnesium combines with one atom of sulphur to form one molecule of magnesium sulphide.

(e) The atoms of magnesium and those of sulphur cannot be further subdivided.

1. According to the assumptions (a) and (b), the atoms of magnesium and sulphur may be represented by the following diagrams (Fig. 4.3).

the atoms in magnesium



the atoms in sulphur

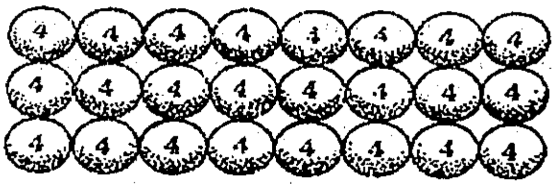


Fig. 4.3 Diagrammatic representation of atoms

2. When magnesium and sulphur react to form magnesium sulphide, each atom of magnesium combines with an atom of sulphur to form a molecule of magnesium sulphide as represented below.

Magnesium atom | sulphur atom | Magnesium sulphide molecule

Now what is the mass ratio of magnesium and sulphur in a molecule of magnesium sulphide?

Since a molecule is formed by the union of a magnesium atom and a sulphur atom, the mass ratio of magnesium to sulphur in the molecule is 3:4.

3. Now we can assume that the molecules of magnesium sulphide are similar with one another. So the structure of magnesium sulphide can be represented as follows :

Since (1) the mass ratio magnesium to sulphur in each molecule of magnesium sulphide is 3:4, and

(2) all the molecules of magnesium sulphide are similar to one another, it is clear that the mass ratio of magnesium to sulphur in the whole compound is also 3:4.

Therefore, the law of definite proportions can be explained by the idea of atoms.

The law of conservation of mass can be explained as follows :

In the combination of magnesium and sulphur, no atoms are destroyed and no new atoms are created. The only change that occurs is the change in the arrangement of atoms. We have seen that each atom of magnesium combines with one atom of sulphur forming a molecule of magnesium sulphide. There is no change in the total number of atoms during the reaction. If there is no change in the total number of atoms, there will be no change in the total mass.

Thus, the mass remains constant in a chemical reaction. This agrees with the statement of the law of conservation of mass.

The idea of atoms described above was first put forward by John Dalton in 1803. We shall now study his atomic theory.

4.6 Dalton's Atomic Theory

John Dalton (1766-1844), an English school teacher, developed the idea of the atomic theory from the experimental evidence of the two laws, namely ;

(1) The law of definite proportions, and
(2) The law of conservation of mass.

These two laws indicate the existence of atoms in all substances even though these atoms cannot be seen.

The basic principles of Dalton's atomic theory are summarized as follows :

(1) Matter is made up of extremely small particles called atoms.
Atoms are indestructible during chemical reactions.

(2) Atoms can neither be created nor destroyed.

(3) Atoms of one particular element have the same mass, shape, size and other properties.

(4) Atoms of different elements have different masses and different properties.

(5) Atoms of different elements combine in simple ratios to form molecules of compounds.

(6) Molecules of the same compound have the same mass, shape and other properties.

(7) Molecules of different compounds have different masses and different properties.

4.7 The Law of Multiple Proportions

Based on this atomic theory, Dalton had predicted the existence of another chemical law. This law is called the Law of Multiple Proportions.

> If two elements combine to form more than one compound, the different weights of one element that combine with a fixed weight of the other element are in a ratio of small whole numbers.

Now, let us look at some experimental data.

Example 1 : Two oxides of carbon

Carbon combines with oxygen forming two oxides. They are carbon dioxide and carbon monoxide.

In an experiment it was found that the weights of carbon and oxygen in the two oxides were as follows :

oxide	carbon	oxygen
carbon dioxide	2.4 g	6.4 g
carbon monoxide	3 g	4 g

Calculation

Calculating the weights of oxygen that combine with a fixed weight of carbon

Take 1 g of carbon for the fixed weight of one element. Calculate the weight of oxygen that combines with 1 g of carbon in each of the oxides.

$$\therefore \quad \left.\begin{array}{l}\text{Wt. of oxygen that}\\ \text{combines with 1g of}\\ \text{carbon in carbon dioxide}\end{array}\right\} = \frac{6.4}{2.4} = 2.583 \text{ g}$$

$\therefore$ Wt. of oxygen that combines with 1g of carbon in carbon monoxide $\left.\right\} = \frac{4}{3} = 1.333$ g

Finding the ratio of the weights of oxygen 2.583 : 1.333

Dividing by the smallest number

$$\frac{2.583}{1.333} : \frac{1.333}{1.333}$$

1.937 : 1

Rounding off to the nearest whole number 2 : 1

2:1 is a small whole number ratio or a simple ratio.

Question : Using the given data in the above example, calculate the ratio of the weights of carbon that combine with a fixed weight of oxygen.

Example 2: Three oxides of chromium

Chromium and oxygen combine to form three oxides. In an experiment the weights of chromium and oxygen in each oxide are as follows :

oxide	chromium	oxygen
1st oxide	5.2 g	1.6 g
2nd oxide	20.8 g	9.6 g
3rd oxide	1.3 g	1.2 g

Calculation

Calculating the weights of oxygen that combine with a fixed weight of chromium

Take 1 g of chromium for the fixed weight of one element and calculate the weights of oxygen that combine with 1 g of chromium.

$\therefore$ Wt. of oxygen that combines with 1g of chromium in 1st oxide $\left.\right\} = \frac{1.6}{5.2} = 0.3077$ g

$\therefore$ Wt. of oxygen that combines with 1g of chromium in 2nd oxide, $\left.\right\} = \frac{9.6}{20.8} = 0.4615$ g

$\therefore$ Wt. of oxygen that combines with 1g of chromium in 3rd oxide $\left.\right\} = \frac{1.2}{1.3} = 0.9232$ g

Finding the ratio of the weights of oxygen in oxides (1), (2), (3)	0.3077 : 0.4615 : 0.9232		
Dividing by the smallest number	$\frac{0.3077}{0.3077}$:	$\frac{0.4615}{0.3077}$:	$\frac{0.9232}{0.3077}$
	1 :	1.4998 :	3.0000
Rounding off to the nearest decimal place	1 :	1.5 :	3
Multiply by 2 to obtain the smallest whole number	2 :	3 :	6

2:3:6 is a small whole number ratio or a simple ratio.

These experimental data support the truth of the law of multiple proportions.

4.8 Explanation of the Law of Multiple Proportions by the Atomic Theory

According to Dalton's Theory, the atoms are indivisible. When two elements, A and B, combine to form more than one compound, an integral number of A atoms will unite with an integral number of B atoms. The possible combinations may be AB, AB_2, AB_3, A_2B, A_2B_3, etc.

Suppose A and B combine to form three compounds, namely AB, AB_2, and AB_3.

If we keep the weight of A in these compounds to be constant, it can be easily seen that the weights of B are in the ratio of 1: 2: 3.

Since this is a simple ratio, the above chemical combinations are in agreement with the law of multiple proportions.

SUMMARY

In this chapter, the laws of chemical combinations such as the law of conservation of mass , the law of definite proportions, Dalton's atomic theory, the law of multiple proportions have been defined with appropriate examples. Worked out examples are also given so that students may understand the application of these laws in chemical combination problems.

Questions and Problems

1. Two elements, A and B, combine to form a compound. In the compound the composition by weight of A and B is 3:4.
   - (a) What is the weight of the element B that combines with 12 g of A?
   - (b) What is the weight of the element A that combines with 16 g of B?
   - (c) What are the weights of A and B in 21 g of the compound?

2. A compound contains 20% of oxygen.

(a) What is the weight of oxygen in 40 g of the compound?

(b) What is the weight of the compound that contains 5 g of oxygen?

(c) What law leads to the reasoning for your calculation?

3 Five experiments were done to find out the proportion by weight of magnesium and oxygen in magnesium oxide.

The following results were obtained in each experiment.

Experiment	magnesium	oxygen
1	1.5 g	1.0 g
2	1.2 g	0.8 g
3	4.8 g	3.2 g
4	2.4 g	1.6 g
5	6.0 g	4.0 g

(a) By using the results of the experiments, determine the proportion by weight of magnesium to oxygen in magnesium oxide.

(b) Which law of chemical combination is in agreement with the above results?

4. Using the proportion by weight of magnesium to oxygen obtained in question (3), answer the followings:

(a) When magnesium is burnt in air, magnesium oxide is formed. If 3 g of magnesium is burnt completely what is the weight of magnesium oxide?

(b) When magnesium is heated in steam, magnesium oxide is formed. After complete combustion of magnesium, 15 g of magnesium oxide is obtained. What is the weight of magnesium used in this experiment?

(c) What law leads to the reasoning for your calculation?

5. (a) Complete combustion of one gram of carbon always results in the formation of exactly 3.667 g of carbon dioxide.

State the law governing this result.

(b) Complete the following statements in accordance with the law governing the results in (a).

In carbon monoxide.

(i) 1 g of carbon combines with 1.33 g of oxygen

(ii) 2 g .... ( ) g ....

(iii) 3 g .... ( ) g ....

(iv) 4 g .... ( ) g ....

6. When 'x' gram of one substance reacts with 'y' gram of another substance, two products are formed. The weights of the two products are 'a' gram and 'b' gram respectively. The weight relation of the above chemical reaction is

$$x + y = a + b$$

What law is in agreement with the above relation of weights?

7. When 'x' gram of a compound is heated, it gives two substances. The weights of the two substances are 'y' gram and 'z' gram respectively. In the light of the law of conservation of mass write an equation to show the relation by weight of the reactants and the products.

8. When mercury (II) oxide is heated it loses weight due to the escape of a gas.
   (a) What are the products obtained by heating mercury (II) oxide?
   (b Why does mercury (II) oxide lose its weight on heating?
   (c) Does the above observation agree with the law of conservation of mass? Give reasons.

9. If a piece of magnesium is burnt in air it gains weight.
   (a) Why does magnesium gain weight on burning?
   (b) Does this observation agree with the law of conservation of mass? Give reasons.

10. Carbon forms two oxides in which the weight of oxygen combines with 1 g of carbon are 1.33 g and 2.66 g respectively. What law is in agreement with this observation? Give reasons.

11. In three oxides of an element the weights of oxygen that combine with a fixed weight of the element are 1. 13 5 g, 2.2 7 g and 3.405 g respectively.
   (a) Find the ratio of the weights of oxygen that combine with a fixed weight of the element.
   (b) What law is in agreement with the above results?

12. A metal forms two chlorides. In one, 1.727 g of the metal combined with 1.03 g of chlorine. In the other, 2.065 g of the metal combined with 2.465 g of chlorine. Show that the formation of these compounds is in agreement with the law of multiple proportions.

13. An element A combines with bromine forming two bromides. It is found that 7 g of one bromide contains 4.2 g of bromine and 13 g of the other bromide contains 9 g of bromine. Show that these values agree with the law of multiple proportions.

14. An element combines with oxygen to form two different compounds. The percentages of the element and oxygen in the two compounds are as follows :

compound	element	oxygen
1st compound	11.1%	88.9%
2nd compound	5.883%	94.117%

Show that the results agree with the law of multiple proportions.

15. An element combines with oxygen to form two oxides which contain 71.42 percent and 55.556 percent of the element respectively. Show that these data agree with the law of multiple proportions.

16. Illustrate the law of multiple proportions using the following data.

	compound	element	oxygen
(a)	1st compound	79.87%	20.13%
	2nd compound	88.81 %	11.19%

	compound	element	oxygen
(b)	1st compound	92.825%	( )
	2nd compound	86.6%	( )

	compound	element	bromine
(c)	1st compound	25.92%	74.08%
	2nd compound	18.92%	81.08%

	compound	element	hydrogen
(d)	1st compound	( )	20%
	2nd compound	( )	7.69%

	compound	element	chlorine
(e)	1st compound	5.264 g	6.699 g
	2nd compound	6.172 g	11.89g

	compound	element	oxygen
(f)	(1) 8.572 g	4.286 g	( )
	(2) 14.3 10 g	5.724 g	( )

17. In an experiment, 1.99 g of copper (II) oxide was found to contain 1.59 g of copper. In another experiment 5 g of oxygen was found to combine with 19.87 g of copper. State the law governing these results and show how the results illustrate the law.

18. It was found by experiment that 5.5 g of an oxide of carbon contained 1.5 g of carbon. Also it was found that 3 g of carbon combined with 4 g of oxygen. State the law governing these results and show how the results illustrate the law.

19. An element forms three oxides. The compositions of the three oxides are :

oxide	element	oxygen
1st oxide	92.83 g	7.17 g
2nd oxide	90.66 g	9.34 g
3rd oxide	86.62 g	13.38 g

Show that the compositions of the three oxides agree with the law of multiple proportions.

20. The compositions of three oxides of nitrogen are as follows :
15 g of the first oxide contains 7 g of nitrogen.
15.2 g of the second oxide contains 5.6 g of nitrogen.
6.75 g of the third oxide contains 1.75 g of nitrogen.
Show that the data are in accordance with the law of multiple proportions.

21. An element forms three compounds in which the percentages of the elements are 51.13, 25.86 and 20.73, respectively. Show that these values illustrate the law of multiple proportions.

22. Write TRUE or FALSE for the following statements.,

(a) The law of conservation of mass states that matter can neither be created nor destroyed.

(b) The law of definite proportion states that the same compound always contains the different elements in the different proportion by weight.

(c) Atoms can be created and destroyed.

(d) Atoms of different elements have different masses and different properties.

(e) Molecules of the same compound have the same mass, shape and other properties.

(f) The atoms of magnesium can be further subdivided.

(g) The total mass of the substances before and after a reaction is always the same.

23. Fill in the blanks with a suitable word or words.,

(a) Atoms can neither be .... nor destroyed.

(b) Matter is made up of extremely small particles called ....

(c) One atom of magnesium combines with one atom of .... to form one molecule of magnesium sulphide.

(d) The same compound always contains the same elements in the same .... by weight.

(e). If two elements combine to form more than one compound, the different weights of the element that combine with a fixed weight of the other element are in a .... of small whole numbers.

(f) Dalton's atomic theory is based on .... laws.

(g) Atoms of different elements combine in ...... to form molecules of compounds.

(h) Magnesium reacts with .... to form magnesium sulphide.

24. Select the correct word or words given in the brackets.

(a) Atoms are indestructible during, (biological, chemical, physical) reactions.

(b) Atoms of different elements have (same, definite, different) masses and different properties.

(c) The atoms of magnesium and those of sulphur (can, cannot, may) be further subdivided.

(d) The mass (ratio, weight, proportion) of magnesium to sulphur in each molecule of magnesium sulphide is 3:4.

(e) Molecules of the (same, different, some) compounds have the same mass, shape and other properties.

25. Match each of the items given in list A with the appropriate item given in list B.

List A	List B
(a) Carbon combines with oxygen forming two oxides	(i) the law of definite proportion and the law of conservation of mass
(b) Magnesium and sulphur on heating give the compound	(ii) the same compound always contains the same elements in the same proportions by weight
(c) Dalton's atomic theory is based on two laws	(iii) carbon dioxide and carbon monoxide
(d) The law of definite proportions	(iv) the law of conservation of mass
(e) Matter can neither be created nor destroyed	(v) magnesium sulphide

26. Answer the following questions.

(a) Write down any four assumptions of Dalton's atomic theory.
(b) Define the law of multiple proportions.
(c) Are atoms of the same element are of the same kind ?
(d) Are the properties of iron atoms same or different ?
(e) Are the properties of a molecule same or different from its component atoms ?
(f) What kind of atoms are present in the same element ?
(g) Why do different elements have different properties ?
(h) Why does the same element have the same properties?

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CHAPTER 5

ARRANGEMENT OF ATOMS AND MOLECULES IN SUBSTANCES

In Chapter 4 we have studied some experimental evidence which made us believe that the substances are composed of atoms and molecules.

In this chapter we shall discuss how atoms and molecules are arranged in the substances.

5.1 Individual Atoms and Groups of Atoms

Atoms

The basic unit of an element that can enter into chemical combination. Atom consists of a small dense nucleus of protons and netrons surrounded by moving electrons.

Individual atoms

The atoms of inert gases, helium, neon, argon, krypton, xenon, exist singly in nature and they do not combine with one another.

Groups of atoms

Excepting the inert gases, the atoms of other elements do not exist singly under normal conditions. They combine with one another to exist as groups of atoms.

5.2 Different Types of Structure

At this stage it is convenient to classify the types of structure according to the composition of the substance, whether it is made up of molecules or of atoms. On this basis, there are two types of structure, the molecular and non-molecular.

The substances which are made up of molecules are molecular substances. The structures of these substances are called molecular structures.

The substances which are not made up of molecules, but are instead, made up of atoms, are called non-molecular substances.

5.3 Molecules and Molecular Structures

Molecules

A molecule is a group of atoms joined together. It is the smallest particle of a substance which can exist by itself.

Molecules in elements and compounds

Molecules exist in elements as well as in compounds. A molecule of an element consists of atoms of the same kind. A molecule of a compound consists of more than one kind of atoms. The atoms of different elements in the molecule of a compound are combined in a definite ratio.

Molecules and three physical states of matter

Gases are made up of molecules each of which can move freely in all directions.

Practically every liquid that occurs at room temperature is a molecular substance; i.e, it is made up of molecules of the substance.

Some solids are molecular substances because they are composed of molecules. But some solids are non-molecular because they are not made up of integral molecules. Instead, they are made up of atoms, which are arranged together in a specific and regular manner to form the whole structure of the solid. This structure is called giant structure.

5.4 Giant structure

The giant structure of a solid is an assembly of atoms arranged in a three dimensional pattern that repeats itself regularly to form the whole mass of the solid.

Diamond is a good example representing the giant structure of atoms.

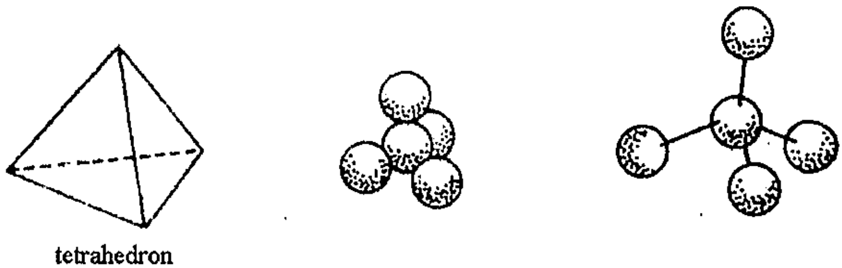


Fig 5.1 Tetrahedral arrangement of carbon atoms

Examine the structure of diamond. Diamond is composed of carbon atoms. Five carbon atoms are arranged in three dimensional pattern as shown in Fig. 5.1. This arrangement is known as a tetrahedral arrangement. This pattern repeats itself regularly to form the structure of the whole mass of diamond as shown in Fig. 5.2.

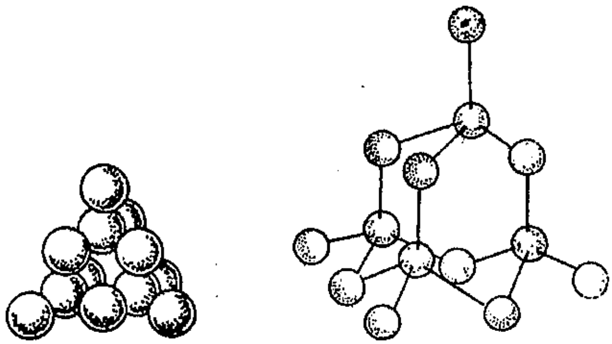


Fig. 5.2 Structure of diamond

5.5 Crystals and Crystalline Substances

Crystals

Crystals are solids which are bounded by plane surfaces at definite angles. Each type of crystal has a regular shape in the form of a polyhedron. (A polyhedron is a solid figure bounded by plane surfaces.) The external regularity of the crystal is due to a regular arrangement of atoms or molecules in the crystals.

Crystalline substances

The regular arrangement of atoms or molecules in the solid gives it a characteristic crystalline shape. The solids with crystalline shapes are called crystalline substances.

5.6 Arrangement of Particles in Gases, Liquids and Solids

The way particles are arranged in a substance depends on whether the substance is in the gaseous, liquid or solid state.

Molecules in gases

Gases are composed of molecules which are in constant and rapid motion.

The gas molecules move freely in all directions until they hit the walls of the container. There exist some weak attractive forces, called van der Waals forces, between gaseous molecules. The gas in a container spreads out to occupy the whole space of the container taking its shape and volume. Hence, gases do not have a definite volume and shape of their own.

A molecule of a gaseous element consists of two or more atoms of the same kind. A molecule of a gaseous compound consists of more than one kind of atom.

Representation of the molecules of some gaseous elements

A sphere represents one atom of an element.

(1) A molecule of oxygen gas may be represented as
How many oxygen atoms are present in an oxygen molecule?

(2) A molecule of nitrogen gas may be represented as
How many nitrogen atoms are present in a nitrogen molecule?

(3) A molecule of hydrogen gas may be represented as
How many hydrogen atoms are present in a hydrogen molecule?

(4) A molecule of chlorine gas may be represented as
How many chlorine atoms are present in a chlorine molecule?

(5) A molecule of bromine vapour may be represented as
How many bromine atoms are present in a bromine molecule?

(6) A molecule of iodine vapour may be represented as
How many iodine atoms are present in an iodine molecule?

(7) A molecule of ozone gas may be shown as
How many oxygen atoms are present in an ozone molecule?

Representation of the molecules of some gaseous compounds

 = Carbon atom

 = Oxygen atom

 = Hydrogen Atom

(1) A molecule of carbon dioxide may be represented as 
How many atoms of carbon and oxygen are present in a molecule of carbon dioxide?

(2) A molecule of carbon monoxide may be represented as 
How many atoms of carbon and oxygen are present in a molecule of carbon monoxide?

(3) A water molecule may be represented as 
How many atoms of hydrogen and oxygen are present in a molecule of water?

In Chapter 7 we shall study the molecular formula that shows the numbers of atoms of elements present in a molecule of a compound.

Molecules in liquids

Liquids are composed of molecules which are not fixed in any definite positions as in the solids. They are able to move freely but not independently as the gas molecules.

The molecules of a liquid readily flow under the influence of gravity so that the liquid can take up the shape of the container.

Liquids can have a definite volume, because the molecules in a liquid are held together more strongly by the van der Waals forces than the molecules in gases.

Bromine is an element which exists in the liquid state at room temperature.

Water is a compound which exits in the liquid state between 0°C and 100°C under atmospheric pressure.

Molecules in solids

Some solids are composed of molecules. These are molecular solids. Some solids are composed of atoms arranged in giant structure.

The particles (atoms or molecules) in a solid are packed very tightly together and held at fixed positions.

For this reason, solids retain their definite shape and definite volume.

Molecular structure of solid elements : Some of the solid elements are constituted of molecules that consist of the same kind of atoms.

Solid iodine is an example of a molecular solid element. Each molecule of iodine consists of two atoms of iodine. The molecules of iodine are arranged regularly so that solid iodine has a crystalline form.

Molecular structure of solid compounds : Most of the solid compounds are constituted of molecules of that compound. These molecules consist of atoms of at least two different elements.

Ice and naphthalene (moth ball) are examples of molecular solid compounds.

Elements having giant structure

Non-metallic elements : Diamond is an element. It is made up of carbon atoms. The structure of diamond has already been discussed.

Metallic elements : At first sight, it is not obvious that metals are crystalline because they have no appearance like diamond or copper (II) sulphate crystals.

On closer inspection they are composed of small crystals known as grains. The arrangement of metal atoms in each grain is in a regular pattern. But the arrangement of these grains in the whole mass of the metal does not follow a regular pattern.

Though metals in bulk show no crystalline appearance, they are in fact, composed of crystals. These crystals have the giant structure built up from the same kind of atoms.

Compounds having giant structure : Compounds which we classify as salts have no separate molecules at all. Thus, they are non-molecular substances. They are compounds with giant structure. In these compounds the atoms of different elements are arranged in a three dimensional pattern that repeats itself regularly to form the whole structure of the compounds.

We shall use common salt as an example. Common salt is a compound of two elements sodium and chlorine. Sodium atoms combine with chlorine atoms, forming a cubic arrangement of atoms (Fig. 5.3).

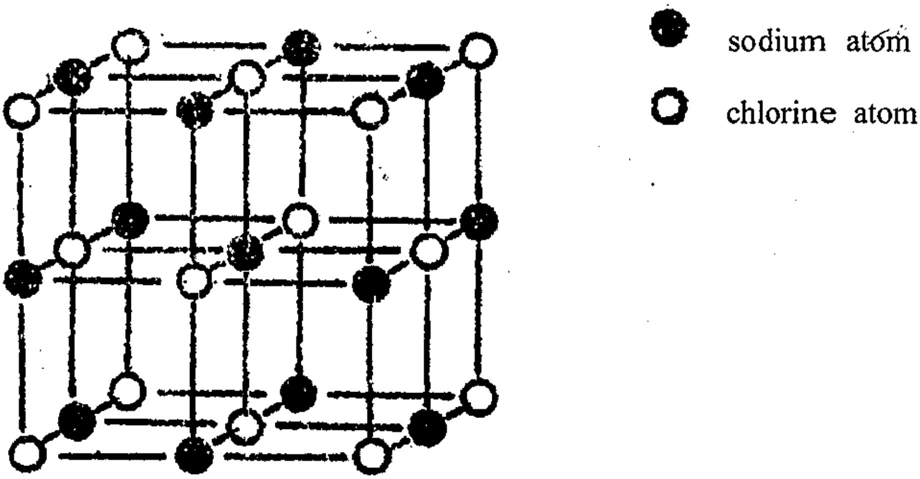


Fig.5.3 Structure of sodium chloride

To test this ability to distinguish between molecular and non-molecular substances, the students should try answering the following questions.

(1) Of what are common gases composed?
(2) Does a gas have a giant structure?
(3) What are the particles present in a substance, which exists in the gaseous state (i) if it is an inert gas (ii) if it is an element other than inert gas (iii) if it is a compound?
(4) Of what is a liquid compound composed?
(5) Does a liquid have a giant structure?
(6) What particles are present in a substance, either an element or a compound, which exists in the liquid state?
Iodine sublimes easily on heating. Decide whether iodine changes easily from the solid state to the gaseous state.
(7) Is solid iodine a molecular substance or a non-molecular substance?

If you can answer the above questions correctly, you know already how to distinguish between molecular substances and non molecular substances. If not, remember the following points.

(1) If a substance is in a gaseous state at room temperature, it may be a molecular substance.
(2) If a substance is in a liquid state at room temperature, it may be a molecular substance.
(3) If a solid can be changed easily from one physical state to another, that is, if it has low melting point and low boiling point, the substance may be molecular.

(4) If a solid cannot be changed easily from one physical state to another, that is, if it has high melting point and high boiling point, the substance may be non-molecular.

(5) Metals and salts are non-molecular substances. They have giant structures.

5.7 Explanation of Changes of State in terms of Energy of Particles

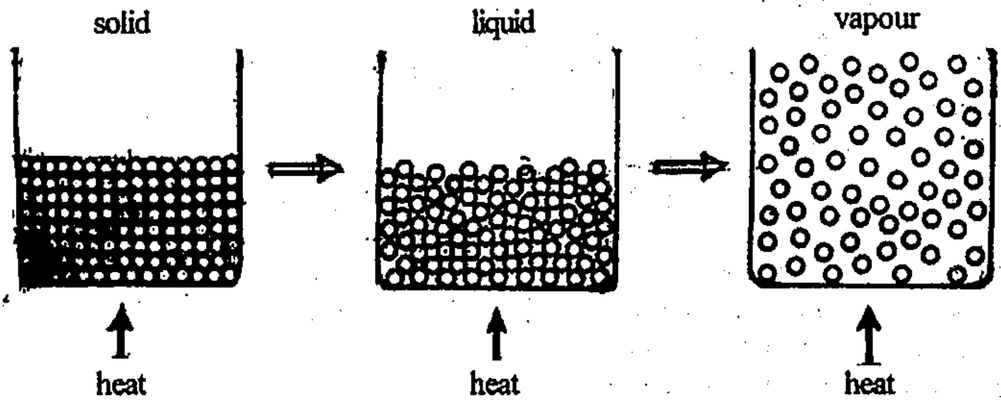


Fig. 5.4 Representation of change of state

When a solid substance is heated it melts. The melting of the solid can be explained in terms of particles in the following way.

As the solid is heated, the heat energy make the particles of the solid vibrate more and more vigorously. Finally these particles gain just enough energy to move out of their fixed positions, but are not yet able to break away from one another.

When all the particles in the solid can move out of the fixed positions, the melting is complete. At that point the solid completely changes into a liquid.

When heating is continued, the particles of the liquid gain more energy and are able to break away from one another. Then the molecules escape from the surface to the space above the liquid. After that, the particles in the vapour or gaseous state move freely in all directions. These free particles constitute the vapour of the substance. In this way a liquid changes into a vapour or a gas (Fig. 5.4).

SUMMARY

This chapter is concerned with the arrangement of atoms and molecules in substances.The terms such as molecular substances and non-molecular substances are well defined. The existence of molecules in elements and compounds have been explained. The giant structure of a solid is exemplified by the structure of diamond. Crystals and crystalline substances have been elaborated. The three physical states of

matter such as solid, liquid and gaseous states have been correlated with molecular nature of the substances. Elements having giant structures are classified into non-metallic elements and metallic elements.

Questions and Problems

1. Do the following normally exist as molecular substances or as non-molecular substances?
   (a) Hydrogen (gas)
   (b) Iron
   (c) Graphite (solid with high melting point)
   (d) Carbon dioxide (gas)
   (e) Aluminium
   (f) Calcium chloride (salt)
   (g) Sulphur (solid, easy to melt)
   (h) Copper (II) sulphate (salt)
2. The melting points and the boiling points of some substances are given below. Pick out the substances having
   (a) molecular structure.
   (b) non-molecular structure.

Substance	Melting point °C	Boiling point °C
Ethanol	-117	78.5
Ammonia	-77.7	-33.4
Magnesium chloride	708	1412
Copper	1083	2582
Carbon tetrachloride	- 23	76.8
Sodium chloride,	801	1420

3. Four expressions are given below.
   (1) Molecular substance (s).
   (2) Non-molecular substance (s).
   (3) Low melting point (s) and boiling point (s).
   (4) High melting point (s) and boiling point (s).

   Fill up the blanks with the most suitable expression mentioned above.
   (a) Common gases at ordinary temperature are ....
   (b) All liquids at ordinary temperature are ....
   (c) Sulphur dioxide is a gas. So it is a ....
   (d) A compound has a low boiling point and a low melting point. Therefore the compound is ....
   (e) All metals are .... Therefore, they have ....
   (f) All salts have .... Therefore, we can decide that salts are ....

(g) An element has .... Therefore, we can decide that the element is ....

(h) A compound which exists in the liquid state at ordinary temperature is a ....

4. A piece of ice is heated to change into three physical states successively. Explain the changes of state in terms of the arrangement of water molecules.
5. Write on the sublimation of solid iodine as a molecular phenomena.
6. Write TRUE or FALSE for each of the following statements.
   (a) The atoms of inert gases helium, neon, argon, krypton, xenon, exist singly in nature.
   (b) The substances which are made up of atoms are molecular substances.
   (c) A molecule of an element consists of atoms of the same kind.
   (d) Bromine is an element which exists in the solid state at room temperature.
   (e) Gases are made up of molecules each of which can move freely in all directions
7. Fill in the blanks with a suitable word or words.
   (a) A molecule is a group of ....joined together.
   (b) Liquids are composed of .... which are not fixed in any definite position as in the solid.
   (c) Ice and naphthalene (moth ball) are examples of molecular .... compounds.
   (d) Diamond is a good example representing the .... structure of atoms.
   (e) The solids with crystalline shapes are called .... substances.
8. Select the correct word or words given in the brackets.
   (a) The substances which are made up of molecules are (atomic, molecular, non-molecular) substances.
   (b) Gases are composed of (particles, atoms, molecules) which are in constant and rapid motion.
   (c) Solid iodine is an example of a (atomic, molecular, non-molecular) solid element.
   (d) Metals and salts are (molecular, non-molecular, elemental) substances, which have gaint structure.
   (e) Bromine is an element which exists in the (gaseous, liquid, solid) state at room temperature.
   (f) A molecular substance can exist as (molecule, giant structure, gas in general).

9. Match each of the items given in list A with the appropriate item given in list B

	List A		List B
(a)	Diamond is a good example	(i)	which are not fixed in any definite position as in the solids.
(b)	Gases are composed of molecules	(ii)	representing the giant structure of atoms
(c)	Liquids are composed of molecules	(iii)	sodium and chlorine
(d)	Solid iodine is an example molecular element	(iv)	which are in constant and of rapid motion
(e)	Common salt is a compound of two elements	(v)	which has two atoms of iodine

10 Answer the following questions.

(a) What are atoms ?

(b) Explain the terms 'molecules' and three physical states of matter.

(c) Is solid iodine a molecular solid or a solid of giant structure ?

(d) Is there any gas which has a giant structure ?

(e) Classify the following substances under two groups:

(A) molecular substance (B) non-molecular substance

(i) ethanol (ii) magnesium chloride (iii) iron (iv) sulphur

(v) copper (vi) ammonia

(f) What is the giant structure? Name metallic elements having giant structure and compounds having giant structures.

****

CHAPTER 6

MASSES OF ATOMS AND MOLECULES

6.1 The Magnitude of the Masses of Atoms

All substances are made up of atoms which have definite masses. But atoms are very small. Consequently, their masses will also be very small.

For example :

the mass of 1 atom of hydrogen	=	1.6734×10^{-24}g
the mass of 1 atom of oxygen	=	2.656×10^{-23}g
the mass. of 1 atom of carbon	=	1.994×10^{-23}g

If we were to express the masses of atoms in grams, the numbers will be so small that we shall need more than twenty decimal places to express them.

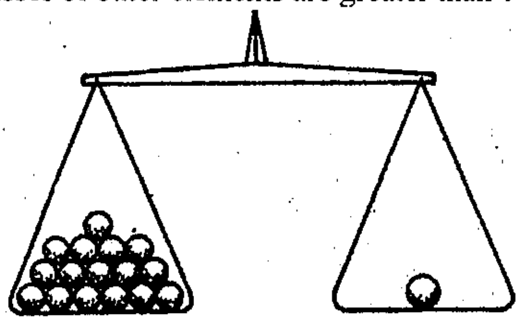
Therefore, the gram unit is too large for expressing atomic masses. It will be very inconvenient for chemical calculations or other practical usages. Besides, the relation between atomic masses cannot be clearly seen. As an example, we cannot see immediately that an atom of oxygen weighs 16 times as much as a hydrogen atom.

Thus, we need a more convenient unit for expressing atomic masses.

A logical unit would be the mass of a certain selected atom itself. It would be most convenient to measure atomic masses by comparison with the mass of such as selected atom as a standard.

6.2 The Hydrogen Scale of Atomic Mass

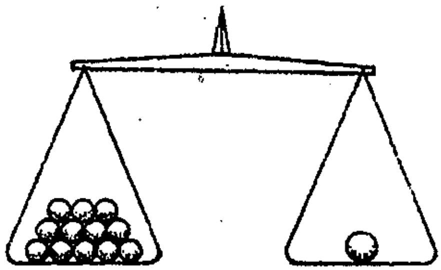
As hydrogen is the lightest element, an atom of hydrogen is the lightest of all atoms. In the hydrogen scale the mass of an atom of hydrogen is taken as the basic unit. This unit is called 1 amu*. The masses of all other atoms are measured in terms of this unit, so that the atomic masses of other elements are greater than 1 amu.



16 atoms of hydrogen 1 atom of oxygen

*amu = atomic mass unit

The mass of an oxygen atom is practically 16 times heavier than that of an atom of hydrogen. Therefore, the atomic mass of oxygen is 16 amu on this standard.



12 atoms of hydrogen 1 atom of carbon

The mass of a carbon atom is 12 times the mass of a hydrogen atom. Therefore, the atomic mass of carbon is 12 amu on this standard.

The atomic mass of magnesium is 24 amu. How many times is an atom of magnesium heavier than a hydrogen atom?

6.3 Reasons for Discarding the Hydrogen Atom as Standard

By taking the mass of the lightest atom (hydrogen atom) as the standard unit, the atomic masses of all other elements are greater than 1 amu. But one draw-back of the hydrogen scale is that hydrogen cannot directly combine with all elements to form compounds.

It is known that oxygen combines directly with most other elements to form compounds. So it is easier to compare the masses of other atoms with the mass of an oxygen atom. For this reason hydrogen was later replaced by oxygen as the standard.

6.4 The Oxygen Scale of Atomic Mass

In the oxygen scale the mass of an atom of oxygen is assigned exactly 16 amu.

$$\text{Therefore 1 amu} = \frac{1}{16} \times \text{mass of an oxygen atom}$$

The masses of other atoms are expressed in terms of this standard unit. When the atomic mass of oxygen is taken as 16 amu, the mass of a hydrogen atom is slightly greater than 1 unit. It is 1.008 amu.

The atomic mass of carbon on this scale, is 12.01 amu.

6.5 The Carbon Twelve (^{12}C)Scale of Atomic Mass

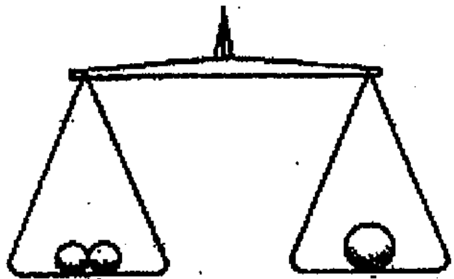
Before 1919 the atomic masses were found the combining weights of the elements obtained by chemical analysis. By 1919 mass spectrometer could be used to measure atomic masses more accurately. In 1961 scientists agreed to choose carbon twelve (^{12}C) as the standard. One reason is that ^{12}C is the most convenient standard in determining atomic masses by a mass spectrometer. In the carbon twelve scale an atom of ^{12}C is assigned a mass of 12 amu.

$$\text{Therefore 1 amu} = \frac{1}{12} \times \text{mass of a } ^{12}C \text{ atom}$$

The masses of other atoms are expressed in terms of this standard unit.

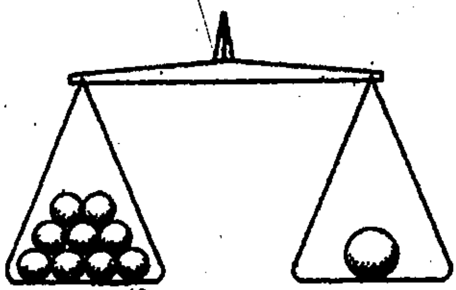
The unit for determining atomic masses is defined as one twelfth the mass of an atom of ^{12}C.

The atomic mass of sulphur is 32 amu. How many times is a sulphur atom heavier than $\frac{1}{12}$ of a ^{12}C atom?



2 atoms of ^{12}C 1 atom of magnesium

A magnesium atom is 2 times heavier than a ^{12}C atom. What is the atomic mass of magnesium?,



9 atom of ^{12}C, 1 atom of sliver

A silver atom is 9 times heavier than a ^{12}C atom. What is the atomic mass of silver?

6.6 Relative Atomic Masses

The atomic masses expressed in amu are really relative atomic masses. The atomic masses expressed in this unit actually represent the numbers of times they are the heavier than the standard unit, i.e., 1 amu.

For illustration, the atomic masses of hydrogen (1.008 amu) oxygen (16 amu), and magnesium (24 amu) are in the ratio of 1.008 : 16: 24 . So, these numbers (1.008, 16, 24) can be regarded as relative atomic masses. They are sometimes also called atomic weight.

Relative atomic masses (i.e., the atomic masses without amu unit) are very useful in calculation. We shall use them wherever convenient.

The relative atomic mass of an element can be defined as the ratio of the mass of an atom of that element to one-twelfth the mass of a carbon twelve atom.

$$\text{Relative atomic mass of an element} = \frac{\text{the mass of an atom of the element}}{\frac{1}{12} \times \text{the mass of a } ^{12}C \text{ atom}}$$

6.7 Atoms of Different Masses in the Same Element

With the help of the mass spectrometer it has been discovered that the same element may consist of a mixture of atoms having different masses.

Isotopes

Atoms of the same element with different masses are called isotopes.

Isotopes of chlorine

Chlorine consists of two isotopes. One isotope has the relative mass 35 and the other isotope has the relative mass 37. The mass of an isotope is called an isotopic mass.

Isotopes of carbon

Natural carbon consists of three isotopes. Their relative masses are 12, 13 and 14, respectively.

On the carbon twelve scale, the standard atom is the carbon isotope of relative mass 12. Therefore the symbol ^{12}C represents the carbon isotope of relative mass 12, which is called carbon twelve.

Isotopes of other elements

Most elements occurring in nature have isotopes. The relative abundances of the isotopes of some common elements are tabulated in Table 6.1.

Table 6.1 Relative abundance of some isotopes

Element	*Isotope*	*Percentage*
Hydrogen	^{1}H	99.98%
	^{2}H	0.02%
	^{3}H	Trace
Oxygen	^{16}O	99.76%
	^{17}O	0.04%
	^{18}O	0.20%
Nitrogen	^{14}N	99.64%
	^{15}N	0.36%
Chlorine	^{35}Cl	75.00%
	^{37}Cl	25.00%
Carbon	^{12}C	98.89%
	^{13}C	1.11%
	^{14}C	Trace
Magnesium	^{24}Mg	78.6%
	^{25}Mg	10.1%
	^{26}Mg	11.3%

6.8 Average Relative Atomic Masses

Since an element consists of atoms of different masses, the mass of any one atom cannot represent the mass of each and every atom in the element.

To overcome this difficulty the average relative atomic mass is used to represent the relative atomic mass of an element.

The average relative atomic mass can be calculated on the basis of the natural abundance of different isotopes.

The relative atomic mass is actually the weighed mean of the relative isotopic masses of an element.

To calculate the relative atomic mass of an element, we must know;

(a) the isotopic masses of the various isotopes present in the element,

(b) the relative abundance of the isotopes in nature.

Example :

Chlorine consists of two isotopes. One isotope has the relative mass 35 and its relative abundance is 75%. The other isotope has the relative mass 37 and its relative abundance is 25%.

Calculation

Isotopic mass	Per cent relative abundance
35.00	75%
37.00	25%

In 100 atoms of chlorine, 75 have the relative mass 35 and 25 have the relative mass 37.

Therefore,

the total relative mass of the 100 atoms = $75 \times 35 + 25 \times 37$

$$= 2625 + 925$$
$$= 3550$$

Average relative mass of 1 atom $= \frac{3550}{100} = 35.5$

The average relative atomic mass of chlorine $= 35.5$

A table showing the average atomic mass of various elements is given in Appendix 2.

6.9 Molecular Masses

The unit used for measuring molecular masses is the same as the unit used for measuring atomic masses.

Hence, we shall describe molecular masses in terms of amu.

The mass of a molecule is obviously the total mass of the atoms present in a molecule.

How can we calculate the relative molecular mass of a molecule?

 = Oxygen atom

 = Carbon atom

 = Carbon dioxide molecule

From this description of the carbon dioxide molecule you should be able to answer the following questions.

- How many carbon atoms are there in a molecule of carbon dioxide?
- How many oxygen atoms are there in a molecule of carbon dioxide?

  The atomic mass of carbon is 12 amu. The relative atomic mass of oxygen is 16 amu.
- What is the mass of the carbon atom in a molecule of carbon dioxide?
- What is the total mass of the oxygen atoms in a molecule of carbon dioxide?
- What is the total mass of the carbon atom and the oxygen atoms in a molecule of carbon dioxide?
- What is the mass of a molecule of carbon dioxide?

  You should now have no difficulty in understanding how we deduce the relative molecular mass of carbon dioxide.

mass of one atom of carbon	+	mass of two atoms of oxygen	=	mass of one molecule of carbon dioxide
12 amu	+	2 (16) amu	=	44 amu

Therefore,

(1) if you know the number of atoms present in a molecule, and

(2) if you know the atomic mass of each atom present in a molecule,

you can easily calculate the molecular mass.

Mass of a molecule	=	The sum of the masses of all atoms present in the molecule

Relative molecular masses

We have mentioned earlier that the atomic mass shows the number of times the atom is heavier than the standard mass, i.e., 1 amu.

Similarly, the molecular mass shows the number of times the molecule is heavier than the standard mass, i.e., 1 amu.

As the atomic masses without amu units are called the relative atomic masses, the molecular masses without amu units are called the relative molecular masses.

The relative molecular mass can be defined as the ratio of the mass of a molecule to one-twelfth the mass of a carbon twelve atom.

$$\text{Relative molecular mass of a molecule} = \frac{\text{The mass of the molecule}}{\frac{1}{12} \times \text{mass of a } ^{12}\text{C atom}}$$

The relative molecular mass of a substance is the sum of the relative atomic masses of atoms present in the molecule.

Example: Carbon dioxide

Relative molecular mass of carbon dioxide	=	Relative mass of one carbon atom	+	Relative mass of two atoms of oxygen
	=	12	+	2 x 16
	=	12	+	32
	=	44		

Example : Relative molecular mass of water

One molecule of water contains one atom of oxygen and two atoms of hydrogen.

Relative molecular mass of water	=	Relative mass of one oxygen atom	+	Relative mass of two atoms of hydrogen
	=	16	+	2 x 1.008
	=	16	+	2.016
	=	18.016		

SUMMARY

All substances are made up of atoms which have definite masses.

In the hydrogen scale the mass of an atom of hydrogen is taken as the basic unit and this unit is called 1 amu.The mass of an oxygen atom is practically 16 times heavier than that of an atom of hydrogen and therefore the atomic mass of oxygen is 16 amu.

$$\therefore\ 1\text{ amu} = \frac{1}{16} \times \text{mass of an oxygen atom}$$

The unit for determining atomic mass is defined as one twelfth the mass of an atom of ^{12}C. The relative atomic mass of an element can be defined as the ratio of the mass of an atom of that element to one twelfth the mass of a carbon twelve atom.

Atoms of the same element with different masses are called isotopes.

eg. $^{35}_{17}Cl$, $^{37}_{17}Cl$,

The relative molecular mass can be defined as the ratio of the mass of a molecule to one twelfth the mass of a carbon twelve atom. The relative molecular mass of a substance is the sum of the relative atomic masses of atoms present in the molecule.

Questions and Problems

Refer to the table of atomic masses given in Appendix 2.

1. The relative atomic mass of nitrogen is 14. How do you understand this statement?
2. A caesium atom is 133 times heavier than one-twelfth of ^{12}C atom. What is the atomic mass of a caesium atom?
3. How many times are the following atoms heavier than one-twelfth of ^{12}C atom?
(a) Magnesium (b) Sulphur (c) Oxygen
4. How many times are the following atoms heavier than an oxygen atom?
(a)Sulphur (b) Copper
5. How many atoms of oxygen are equal in mass with one atom of copper?
6. A titanium atom is four times heavier than a ^{12}C atom. What is the mass of a titanium atom?
7. If a magnesium atom is chosen as the standard, what relative atomic mass should be assigned to the magnesium atom so that atomic mass of the lightest element hydrogen is about 1?

8. A molecule of carbon dioxide consists of one atom of carbon and two atoms of oxygen.
   (a) What is the molecular mass of carbon dioxide?
   (b) How many times is a carbon dioxide molecule heavier than one-twelfth of ^{12}C atom?

9. One atom of sulphur and two atoms of oxygen combine to form a molecule of sulphur dioxide. A sulphur dioxide molecule is sixteen times heavier than one atom of helium. What is the relative atomic mass of helium?

10. In a molecule of a compound there are two atoms of hydrogen and two atoms of another element. The relative mass of the molecule is 34. What is the relative atomic mass of the other element?

11. In naturally occurring lead there are four kinds of atoms with the following relative masses.
   (a) 204 (b) 206 (c) 207 (d) 208
   How many isotopes are there in naturally occurring lead?

12. The isotopic masses of naturally occurring oxygen and their relative abundance are as follows:

Isotopic mass	Per cent relative abundance
16	99.76%
17	0.04%
18	0.2%

   What is the average relative atomic mass of oxygen?

13. Explain the following statement.
   The average relative atomic mass of an element is the weighed mean of the masses of all atoms present in the element.

14. Write a short note on the different scales used for measuring atomic masses.

15. Write TRUE or FALSE for the following statements.
   (a) All substances are made up of atoms which have definite masses.
   (b) The mass of an oxygen atom is practically 16 times lighter than that of an atom of hydrogen.
   (c) The mass of a carbon atom is 12 times the mass of a hydrogen atom.
   (d) One draw-back of the hydrogen scale is that hydrogen cannot directly combine with all elements to form compounds.
   (e) Atoms of the different elements with different masses are called isotopes.
   (f) The mass of ^{12}C atom is 12 g.
   (g) The mass of a carbon atom is 12 times the mass of an oxygen atom.
   (h) The relative atomic mass of oxygen is 16 amu.

16. Fill in the blanks with a suitable word or words.
   (a) Atoms of the same element with different masses are called ....

(b) The mass of an oxygen atom is practically 16 times .... than that of an atom of hydrogen.

(c) An atom of hydrogen is the ....of all atoms.

(d) The atomic mass of ....on this scale is 12.01 amu.

(e) .... is the standard unit of atomic mass.

(f) The molecular mass of water is ....

(g) A copper atom is 4 times heavier than the mass of an oxygen atom. Therefore, the atomic mass of copper is ....

.7. Select the correct word or words given in the bracket.

(a) The mass of a carbon atom is 12 times (heavier, lighter, equal) than to that of the mass of a hydrogen atom.

(b) In the (carbon, hydrogen, oxygen) scale the mass of an atom of oxygen is assigned exactly 16 amu.

(c) The relative atomic mass of an element can be defined as the ratio of the mass of an atom of that element to one-twelfth the mass of a (carbon, hydrogen, oxygen) twelve atom.

(d) Atoms of the same element with (same, different, equal) masses are called isotopes.

(e) Among the three atomic mass scales (carbon, hydrogen. oxygen) scale is discarded because it cannot directly combine with all elements to form compounds.

18. Match each of the items given in list A with the appropriate item given in list B.

	List A		List B
(a)	The unit for hydrogen scale of atomic mass	(i)	Atoms of the same element with different masses
(b)	The reason for discarding the hydrogen scale	(ii)	1 amu
(c)	The mass of oxygen atom	(iii)	16 times heavier than that of an atom of hydrogen
(d)	The mass of carbon atom	(iv)	hydrogen cannot directly combine with all elements to form compounds
(e)	Isotopes	(v)	12 times the mass of a hydrogen atom

19. Answer the folowing questions.

(a) The mass of a gas molecule is equivalent to 4 times of ^{12}C . What is the relative molecular mass of the gas ?

(b) An atom of an element weighs 4 times as heavy as 1/12 mass of an atom of ^{12}C. What is the mass of that atom?

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CHAPTER 7

SYMBOLS, FORMULAE AND EQUATIONS

Symbols, Formulae and Equations are used in chemistry for the following reasons.

(1) to save time in describing chemical reactions.

(2) to clearly illustrate how the atoms and molecules react with one another, and

(3) to precisely express the weight relations between the reactants and the products.

7.1 Symbols

What does a symbol mean?

A symbol is a short notation of an element. The first letter in the name of an element is usually chosen as the symbol of the element.

Example:	**H**	is the symbol of Hydrogen
	C	is the symbol of Carbon
	O	is the symbol of Oxygen

When the names of two or more elements begin with the same letter, the first letter together with another letter in the name of an element is chosen as the symbol of the element.

Example:	Ca	is the symbol of Calcium.
	Cl	is the symbol of Chlorine
	Cr	is the symbol of Chromium

A symbol represents an atom of the element.

Example:	H	represents an atom of Hydrogen
	O	represents an atom of Oxygen
	Ca	represents an atom of Calcium

Moreover, a symbol also represents the mass of an atom.

Example:	H	represents the mass of a hydrogen atom
	O	represents the mass of an oxygen atom
	Ca	represents the mass of a calcium atom

So we can see that a symbol conveys the following:

A symbol represents the name of an element, an atom of the element, and the atomic mass of the element.

Symbols of some elements

The symbols of some elements are taken from English names. The symbols of some elements, which have been known since earlier times, are taken from Latin names.

In the following table the Latin names, from which the symbols were taken, are described along with the English names.

In the symbols of some metallic and non-metallic elements are shown in Table 7.1 and 7.2. The Latin names, from which the symbols were taken are described along with the English names.

Table 7.1 Symbols of metallic elements

English Name	*Latin Name*	*Symbol*
Sodium	Natrium	Na
Potassium	Kalium	K
Calcium		Ca
Barium		Ba
Magnesium		Mg
Aluminium		Al
Manganese		Mn
Zinc		Zn
Chromium		Cr
Iron	Ferrum	Fe
Tin	Stannum	Sn
Lead	Plumbum	Pb
Copper	Cuprum	Cu
Mercury	Hydrargyrum	Hg
Silver	Argentum	Ag
Gold	Aurum	Au
Platinum		Pt
Antimony	Stibium	Sb

Table 7.2 Symbols of some non-metallic elements

Name	*Symbol*
Carbon	C
Sulphur	S
Phosphorus	P
Oxygen	O
Hydrogen	H
Nitrogen	N
Chlorine	Cl
Bromine	Br
Iodine	I

7.2 Formulae

What does a formula mean?

Molecular formulae of elements

A molecular formula is a short notation representing a molecule of a substance which may be an element or a compound. It shows the number of atoms contained in a molecule. The elements such as hydrogen, nitrogen, oxygen, chlorine and bromine can exist in the gaseous state. These gaseous elements are composed of molecules containing two atoms each as you have seen in Chapter 5 on page 60.

Thus,

the formula of hydrogen is written as H_2
the formula of nitrogen is written as N_2
the formula of oxygen is written as O_2

What will be the molecular formulae of chlorine and bromine?

The subscript 2 in the above formulae shows that the two atoms in each formula exist in combination.

If we write 2H, it will mean two atoms of hydrogen which are not combined with each other.

A molecular formula also represents a molecule of the corresponding element.

Example:

H_2 means a molecule of hydrogen
N_2 means a molecule of nitrogen
$2O_2$ means two molecules of oxygen

A molecular formula also represents the molecular mass of the corresponding element.

Example

H_2 represents a molecular mass of 2 x 1.008 = 2.016 amu
O_2 represents a molecular mass of 2 x 16 = 32 amu

Molecular formulae of compounds

For those compounds which exist in the form of molecules, a formula represents a molecule as well as the molecular mass of the compound.

Example:

CO_2 represents a molecule of carbon dioxide as well as a molecular mass of 44 amu.

H_2O represents a molecule of water as well as molecular mass of 18.016 amu.

If a molecule of carbon monoxide contains an atom of carbon and an atom of oxygen, what will be the molecular formula of carbon monoxide?

If a molecule of ammonia consists of an atom of nitrogen and three atoms of hydrogen, what will be the molecular formula of ammonia?

If a molecule of nitric acid contains an atom of hydrogen, an atom of nitrogen and three atoms of oxygen, what will be the molecular formula of nitric acid?

Formulae for non-molecular compounds

For those compounds which exist in the form of giant structure a formula represents the simplest unit of the compound.

Example :

In sodium chloride the simplest unit consists of an atom of sodium and an atom of chlorine, because sodium and chlorine combine in the atomic ratio 1 : 1.

So the formula of sodium chloride is Na C1.

Similarly the formula of magnesium oxide is MgO.

If potassium and bromine combine in the atomic ratio of 1:1 to form potassium bromide, what will be the formula of potassium bromide?

If sodium and oxygen combine in the atomic ratio of 2:1 to form sodium oxide, what will be the formula of sodium oxide?

In the above compounds, each formula also represents the formula mass of the corresponding compound. The formula mass is the sum of the masses of atoms present in a formula.

Example :

NaCl represents a formula mass of 23 + 35.5	=	58.5 amu
MgO represents a formula mass of 24 + 16	=	40 amu

We can now sum up the significance of the formula as:

A formula	=	a molecule of an element, in which case it is usually called a molecular formula.
	=	a molecule of a molecular compound and may also be referred to as the molecular formula.
	=	the simplest unit of a non-molecular compound (i.e., compound of giant structure).
	=	the molecular mass or formula mass of the corresponding substance.

Note As compounds of giant structure do not have separate molecules, they cannot have molecular mass. So formula mass is used instead of molecular mass.

7.3 Determination of the Chemical Formulae

To determine the chemical formula of a substance, the following steps are necessary.

1. Determination of the composition by weight of each constituent element in the compound using appropriate analytical methods.

2. Deduction of the empirical formula from the composition by weight.

> **The empirical formula** is the formula representing the relative number of atoms of each kind present in a molecule of a given compound.

The empirical formula is easily obtained by dividing the weight proportion of each element in the compound by the atomic mass of the element.

3. Knowing the molecular mass of the compound, we can use the empirical formula to determine the actual number of atoms of each kind present in the molecule. This will be the molecular formula.

> **The molecular formula of a compound** is the formula which shows the actual number of atoms of each kind present in a molecule of the compound.

Determination of the empirical formula of magnesium oxide as an example

In Chapter 4, on page 45, we have done an experiment to find the weight ratio of magnesium and oxygen in magnesium oxide. Refer to this experiment. We get two weights from this experiment, i.e., weight of magnesium and weight of oxygen which has combined with that weight of magnesium.

Step 1

(a) **Experimental determination of the composition by weight of magnesium oxide**

Wt. of magnesium = 3 g
Wt. of magnesium oxide = 5 g
Wt. of oxygen combining with 3 g of magnesium = (5-3)= 2 g

(b) **Calculation of the proportion by weight of magnesium and oxygen in magnesium oxide**

In magnesium oxide, 3g of magnesium combines with 2 g of oxygen.

∴ The weight ratio of magnesium and oxygen in magnesium oxide = 3 g : 2 g

or 3 : 2

The atomic mass of magnesium is 24
The atomic mass of oxygen is 16
Since the weight ratio of magnesium to oxygen is 3 : 2

Step 2

(a) Calculation of the atomic ratio	the atomic ratio of magnesium to oxygen	$= \frac{3}{24} : \frac{2}{16}$
		$= \frac{1}{8} : \frac{1}{8}$
		$= 1 : 1$
	The simplest ratio of atoms of magnesium and oxygen in magnesium oxide is 1 : 1.	
(b) Writing the empirical formula	∴ The empirical formula of magnesium oxide	$= MgO$

Uses of empirical formula

Magnesium oxide is a compound with giant structure. For compounds of giant structure the molecular weight cannot be found and the empirical formula is used as such.

For compounds of giant structure empirical formulae are used in writing chemical equations.

Determination of molecular formula of hydrogen peroxide as an example

Step 1 (a) The composition by weight of hydrogen peroxide is first determined. It was found by experiment that 1 part by weight of hydrogen combines with 16 parts by weight of oxygen to form hydrogen peroxide.

(b) The weight ratio of hydrogen and oxygen in hydrogen peroxide is then calculated. Since 1 part by weight of hydrogen combines with 16 parts by weight of oxygen, the weight ratio of hydrogen to oxygen is 1:16.

Step 2 (a) To find the atomic ratio, we divide the weight ratio of each element by the atomic mass of the element.

The atomic mass of hydrogen = 1 (nearest)
The atomic mass of oxygen = 16

$$\therefore \text{The atomic ratio} = \frac{1}{1} : \frac{16}{16}$$

$$= 1 : 1$$

(b) The empirical formula of hydrogen peroxide = HO

Step 3 (a) *Determination of the molecular mass of hydrogen peroxide*

The molecular mass of the hydrogen peroxide is determined by experiment. The molecular mass of hydrogen peroxide is found to be 34.

(b) *Determination of the molecular formula from the empirical formula*

From the empirical formula HO, we know that the numbers of atoms of hydrogen and oxygen in hydrogen peroxide is 1: 1. But we do not know yet the actual numbers of atoms of hydrogen and oxygen in a molecule of hydrogen peroxide.

The actual numbers of atoms of hydrogen and oxygen may be 1: 1 or an integral multiple of 1 : 1.

So we can write the molecular formula of hydrogen peroxide as $(HO)_n$ where $n = 1,2,3....$

Since the sum of the atomic mass of all the atoms in a molecule must be equal to molecular mass.

$$(HO)_n = 34$$
$$(1 + 16)_n = 34$$
$$17_n = 34$$
$$n = 2$$
$$(HO)_n = (HO)_2$$
$$= H_2O_2$$

The molecula formula of hydrogen peroxide $= H_2O_2$

Determination of molecular formula of ethyl alcohol as an example

Ethyl alcohol contains 52.18% carbon, 13.04% hydrogen and the rest oxygen. Its molecular mass is 46. What are the empirical and molecular formulae of ethyl alcohol?

To determine the empirical and molecular formulae of ethyl alcohol from the given data, we can follow the same steps as in the previous example. The weight ratios of the constituent elements are:

C = 52.18% ; H = 13.041% ; O = 34.78%

From this, we can find out the atomic ratios :

$$C : H : O$$
$$\frac{52.18}{12} : \frac{13.04}{1.00} : \frac{34.78}{16}$$
or $$4.4 : 13.04 : 2.17$$
or $$\frac{4.4}{2.17} : \frac{13.04}{2.17} : \frac{2.17}{2.17}$$
or $$2.0276 : 6.0092 : 1$$
or $$2 : 6 : 1$$

from which we have the empirical formula of ethyl alcohol as C_2H_6O.

The molecular formula may be written as $(C_2H_6O)_n$.

Then, $(C_2H_6O)_n = 46$

or $(2 \times 12 + 6 \times 1 + 1 \times 16)_n = 46$

or $46 \times n = 46$

$n = 1$

Hence, the molecular formula of ethyl alcohol is C_2H_6O.

Note: The method of determining the molecular formulae of gaseous elements will be illustrated in Chapter 9.

From the examples you have learned in this chapter you will know that the formulae of substances are determined experimentally. Every formula that you will come across in this course is determined experimentally by chemists as you have seen. No formula can be written without experimental determination.

7.4 Equations

What is a chemical equation?

When carbon is burnt in air, carbon combines with oxygen to form carbon dioxide.

This is a chemical change. The process of undergoing a chemical change is called a chemical reaction. A chemical reaction can be represented by a chemical equation. There are two types of chemical equations: i.e.,

- Equation in words
- Equation in symbols

Equation in words

In an equation in words the names of the reactants are written on the left-hand side of the equation and the names of the products on the right-hand side of the equation.

In the burning of carbon, carbon and oxygen are the reactants and carbon dioxide is the product.

So the equation in words will be;

carbon + oxygen = carbon dioxide

Equation in symbols

An equation in symbols describes the reactants and the products in terms of chemical symbols and formulae.

The equation in symbols for the burning of carbon will be ;

$$C + O_2 = CO_2$$

Chemical equations : How they are written

Representation of non-molecular elements

When elements are non-molecular, the symbols are used to represent the elements.

Such elements include all metals and some non-metals like carbon and sulphur.

Example : Na, Zn, Mg, Fe, Cu, C, S

Representation of molecular elements
When elements are molecular such as oxygen hydrogen, etc., molecular formulae are used to represent them.
Example: O_2, H_2, N_2, Cl_2, Br_2, I_2, P_4

Representation of non-molecular compounds
The compounds having giant structures are represented by their empirical formulae.
Example: MgO, $NaCl$, KCl, Na_2O

Representation of molecular compounds
The compounds which exist in the form of separate molecules are represented by their molecular formulae.
Example: CO_2, CO, H_2O, H_2O_2

Steps in writing chemical equations
Let us take the burning of magnesium as an example.

1.	Write the word equation for the reaction.	1. magnesium + oxygen = magnesium oxide
2.	Write the symbols and the formulae of the reactants and the products under the respective names.	2. $Mg + O_2 = MgO$
3.	Balancing the equation in symbols by multiplying the symbols and the formulae with suitable numbers, so that the number of atoms of each element is equal on both side of the equation.	3. $2Mg + O_2 = 2MgO$

In practice, step 2 and 3 are combined as a single step so that the equation becomes

magnesium	+	oxygen	=	magnesium oxide
$2Mg$	+	O_2	=	$2MgO$

(a) Why is the symbol Mg written to represent magnesium?
(b) Why is the formula O_2 written to represent oxygen?

Why is the equation balanced?
According to the Law of Conservation of Mass the total mass of the reactants is equal to the total mass of the products. Therefore the number of the atoms of each element

before and after the reaction must also be equal. For this reason it is necessary to write the balanced equation.

What does a chemical equation express?

A chemical equation describes

(1) the reactants, which are the substances that take part in the reaction
(2) the products, which are the substances that are produced in the reaction
(3) the rearrangement of the atoms during the reaction
(4) the weight relationship of the reactants and the products.

The physical states of the reactants and products

To mention whether the reactants and the products are solid, liquid, gaseous or in solution, the following abbreviations are written after the corresponding symbols and formulae.

$$2Na\,(s) \quad + \quad 2H_2O\,(l) \quad = \quad 2NaOH\,(aq) \quad + H_2\,(g)$$

s	=	solid	aq	=	aqueous or in water solution
l	=	liquid	g	=	gas

The following abbreviations are usually used in chemical equations.

(1) Δ = heat
(2) $\downarrow$ = formation of precipitate
(3) $\uparrow$ = gas evolved
(4) $\rightleftharpoons$ = reversible reaction

Example :

(1) $Ca(HCO_3)_2\,(aq) \xrightarrow{\Delta} CaCO_3(s) + H_2O\,(l) + CO_2\,(g)$

or $Ca(HCO_3)_2\,(aq) \xrightarrow{\Delta} CaCO_3\downarrow + H_2O + CO_2\uparrow$

(2) In enclosed spaces or closed systems some reactions can occur in both directions, that is, the reactants give products which can be reconverted to the reactants. Such reactions are said to be reversible.

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

For a chemical equation to fully convey the essence of the chemical reaction, it should be written as shown in the above examples. In writing chemical equation emphasis is usually placed on ;

(a) writing a balanced chemical equation
(b) mentioning the conditions for the reaction to take place and
(c) disclosing the physical states of the reactants and products.

SUMMARY

This chapter is concerned with symbols, formulae and equations. A symbol is a short notation of an element. It can represent the symbol of an element, it can also represent an atom of the element, it can also represent the mass of an atom, and it can also represent the name of an element. This chapter explains about the symbols of metallic elements and symbols of non-metallic elements. Molecular formulae of elements and molecular formulae of compounds have been differentiated. The concepts of empirical formular and molecular formular of a compound have been clearly defined. The chemical equation can be expressed in words and in symbols. The physical states of the reactants and products must be mentioned in writing chemical equation.

Questions and Problems

1. Describe the derivation of the empirical formula of magnesium oxide under the following headings.
   (a) Experimental details for finding the composition by weight of magnesium and oxygen in magnesium oxide.
   (b) Calculation of the empirical formulae from the weight composition.
2. Calculate the formula of a compound in which 3 g of carbon combines with 4 g of oxygen.
3 Calculate the formula of the following compounds.
   (a) Sodium oxide in which 2.3 g of sodium combines with 0.8 g of oxygen.
   (b) Oxide of phosphorus in which 1.24 g of phosphorus combines with 0.96 g of oxygen.
   (c) Oxide of lead in which 2.07 g of lead combines with 0.32 g of oxygen.
4. Calculate the empirical formula of a nitrogen oxide. It is found that 4.2 g of nitrogen is contained in 16.2 g of the nitrogen oxide.
5. Calculate the formulae of the respective compounds whose weight compositions are given below,
   (a) 1.91 g of oxide of sodium is formed from 1.42g of sodium.
   (b) 2.23 g of oxide of lead is formed from 2.07 g of lead.
   (c) 16.04 g of oxide of iron is formed from 11.22 g of iron.
   (d) 5 g of oxide of copper is formed from 4 g of copper.
6. Calculate the formulae of the following compounds.
   (a) Water in which hydrogen and oxygen combine in the proportion of 1:8 by weight.
   (b) Oxide of carbon in which carbon and oxygen combine in the proportion of 3:8 by weight.
   (c) Oxide of sulphur in which sulphur and oxygen combine in the proportion of 2:3 by weight.

(d) Oxide of iron in which iron and oxygen combine in the proportion of 7:3 by weight.

(e) A compound of calcium and bromine in which the composition by weight of calcium and bromine is in the ratio 1:4.

7. Calculate the formula of an oxide of nitrogen which contains 63.64% by weight of nitrogen.
8. Calculate thc formulae of the following compounds.
   (a) Iron oxide in which the weight of iron is 77.7% and that of oxygen is 22.3%.
   (b) Iron oxide in which the weight of iron is 70%.
   (c) Iron oxide in which the weight of oxygen is 27.6%.
9. In a compound of sulphur and chlorine, 6.4g of sulphur combines with 7.1g of chlorine. Calculate the empirical formula. Then calculate the molecular formula of the compound if the molecular mass of the compound is 135 amu.
10. Find the empirical formulae and the molecular formulae of the following compounds.
   (a) A compound having 92.3% C, 7.7% H and a molecular mass of 78 amu.
   (b) A compound having 26.6% C, 2.23% H, 71.17% O and a molecular mass of 90 amu.
   (c) A compound having 43.7% P, 56.3% O, and a molecular mass of 284amu.
   (d) A compound having 88.2% C, 11.8% H and a molecular mass of 136amu.
11. Calculate the molecular formula of the following compounds.

empirical formula	**molecular mass (amu)**
H	2
O	32
O	48
N	28
Cl	71
Br	160
I	254
SCl	135
NO_2	92
P_2O_5	284

12. You are given the following equation:

$$2Ca + O_2 = 2CaO$$

   (a) What are the reactants?
   (b) What is the product?
   (c) How many atoms of calcium combine with a molecule of oxygen?
   (d) What is the atomic ratio of calcium and oxygen in calcium oxide?

(e) If the atomic mass of calcium is 40 amu and the atomic mass of oxygen is 16 amu, what is the formula mass of CaO?

(f) What is the composition by weight of calcium oxide?

13. Write TRUE or FALSE for the following statements.

(a) Symbols, formulae and equations are used in chemistry to save time in describing chemical reactions.

(b) A symbol is a long notation of an element.

(c) A symbol represents the name of an element, an atom of the element, the atomic mass of the element.

(d) A molecular formula is a long notation representing a molecule of a substance which may be an element or a compound.

(e) The empirical formula is the formula representing the relative number of atoms of each kind present in a molecule of a given compound.

(f) The symbol of Fe, Pb, Cu, Sn, are taken from latin names.

14. Fill in the blanks with a suitable word or words or symbols.

(a) A symbol is a short ....of an element.

(b) Molecular formula shows the number of .... contained in a molecule.

(c) CO_2 represents a molecule of .... as well as a molecular mass of 44.

(d) The formula of magnesium oxide is ....

(e) The symbol for lead is ...... , gold is .... , and mercury is ....

(f) The name of Cr is .... and that of Hg is ....

15. Select the correct word or words given in the bracket.

(a) The name of Na is (calcium, sodium, potassium).

(b) The formula of sodium chloride is (KCl, NaCl, HCl).

(c) The symbol for silver is (Fe, Ag, Cu).

(d) The molecular formula for chlorine is (Cl_2, Br_2 , I_2).

(e) The molecular formula of a compound is the formula which shows the actual number of (group, atoms, elements) of each kind present in a molecule of the compound.

(f) The name of Pt is (Potassium, Phosphorus, Platinum).

16. Match each of the items given in list A with the appropriate item given in list B.

	List A		**List B**
(a)	The symbol of calcium	(i)	MgO
(b)	The molecular formula of nitrogen	(ii)	carbon dioxide is the product
(c)	The formula of magnesium oxide	(iii)	Ca
(d)	In the burning of carbon, carbon and oxygen are reactants	(iv)	equation in words and equation in symbols
(e)	There are two types of chemical equations	(v)	N_2

17. Answer the following questions.

(a) What do the following abbreviations stand for in a chemical equation?

(i) Δ (ii) $\downarrow$ (iii) $\uparrow$ (iv) $\rightleftharpoons$

(b) Give the English names of the following: Mn, Al, K, Sn

(c) What is meant by an empirical formula ?

****

CHAPTER 8

FORMULA WRITING AND THE NAMING SYSTEM

In Chemistry, it is convenient to describe chemical reactions by means of chemical equations, using symbols and formulae. At first sight it seems to be difficult to learn how to write the formulae and the names of a large number of different compounds.

But it is not so difficult if we know

(1) the combining capacity of the atoms of different elements and how to use them in formula writing, and

(2) the rules for naming compounds.

In this chapter we shall discuss how to write the formulae of compounds with the proper use of the combining capacities of the elements. Then we shall learn the rules for naming compounds.

8.1 The Concept of Combining Capacity

Examine the formulae of some compounds shown below.

HCl	$MgCl_2$	$CaCl_2$	$AlCl_3$	CCl_4	H_2O
Na_2O	MgO	CaO	Al_2O_3	CO_2	H_2S

We may notice that, atoms of different elements combine in definite ratios to form compounds. In other words, atoms have definite combining capacities.

We can express this combining capacity by certain numbers.

For example, we can assign 1 to the combining capacity of hydrogen.

Then the combining capacity of chlorine must also be assigned 1, because one atom of hydrogen combines with one atom of chlorine.

In sodium chloride NaCl, one atom of chlorine combines with one atom of sodium. Sodium, therefore, must have a combining capacity of 1.

In H_2O, two atoms of hydrogen combine with one atom of oxygen, which must, therefore, have a combining capacity of 2.

Combining capacity

The combining capacity or valence of an element is represented by the number of atoms of hydrogen, chlorine or sodium that combine with one atom of that element. The term "valence" is also used to express combining capacity.

Different atoms have different combining capacities. Chlorine, for example will combine with sodium, calcium or aluminium to form NaCl, $CaCl_2$ or $AlCl_3$ respectively. From our definition of the combining capacities we can see sodium has a combining capacity of 1, calcium a combining capacity of 2 and aluminium a combining capacity of 3 in these compounds.

Sulphur combines with hydrogen, sodium and calcium to form H_2S, Na_2S and CaS.

In H_2S and Na_2S, sulphur shows a constant combining capacity of 2. Hence the combining capacity of sulphur in its compounds with calcium must also be 2. In this case, the combining capacity of calcium in CaS must be 2. Calcium in $CaCl_2$ also has a combining capacity of 2.

Aluminium in aluminium chloride, $AlCl_3$, shows a combining capacity of 3. Aluminium combines with sulphur to form Al_2S_3. Let us see if it checks with what we already know about the combining capacities of Al and S.

Since each atom of Al has a combining capacity of 3, the total combining capacity of 2Al atoms is 2 x 3 or 6. Sulphur has a combining capacity of 2 in its metallic sulphides. The 3 sulphur atoms in Al_2S_3 must, therefore have a total combining capacity of 3 x 2 or 6.

The total combining capacities of the metallic atoms in Al_2S_3 is exactly balanced by the total combining capacities of the sulphur atoms in it.

In the previous paragraph we have used the idea of the combining capacities of the metallic atoms like Na, Ca and Al being balanced by the combining capacities of the non-metallic atom like Cl and S. Further examples are provided in Table 8.1.

Table 8.1 Formulae of some compounds containing metallic and non-metallic atoms

Metallic element	*Hydride*	*Chloride*	*Oxide*	*Sulphide*
Sodium	NaH	NaCl	Na_2O	Na_2S
Calcium	CaH_2	$CaCl_2$	CaO	CaS
Aluminium	AlH_3	$AlCl_3$	Al_2O_3	Al_2S_3

If we examine chemical compounds, we can see that metals usually combine with non-metals. Metals may not chemically combine with other metals, but non-metals may combine with other non-metals to form compounds. For example, oxygen may combine with nitrogen, sulphur, carbon and so on. You may also examine the metal free compounds like ammonia, NH_3; hydrogen sulphide, H_2S; methane, CH_4; carbon disulphide, CS_2 and carbon tetrachloride, CCl_4.There are many others.

The above observations are readily understood if we accept it as a fact that the elements in a compound have electrical natures which are the opposite of each other. One may be electropositive (tendency to acquire positive electric character) and the other electronegative (tendency to acquire negative electric character). For instance, metals like Na, Ca and Al are supposed to show an electropositive nature in compounds formed by combination with non-metals. It follows, therefore, that the non-metals in these compounds must have electronegative natures.

Similarly in the metal-free compounds like carbon dioxide, CO_2; sulphur dioxide, SO_2 ; nitrogen dioxide, NO_2 ; etc., the oxygen atom shows greater electronegative nature than the other elements which show relatively more electropositive natures. Note that an element like sulphur may show an electronegative nature in combination with metals but an electropositive nature in combination with oxygen.

Do you think you can make similar comments for hydrogen? In which compounds does it show an electronegative nature and in which others does it show an electropositive nature?

Fixed combining capacities of certain elements

The element sodium reacts with other elements to form compounds in which it shows electropositive character and a constant combining capacity of 1.

Other metallic elements like calcium and magnesium always show a combining capacity of 2 in their compounds. Such elements have fixed combining capacities or valencies.

Variable combining capacities of certain elements

Two different compounds of copper and chlorine are known. They are copper (I) chloride CuCl, and copper (II) chloride, $CuCl_2$. Copper also form two oxides copper (I) oxide, Cu_2O, and copper (II) oxide CuO. Obviously, copper in its reactions with other elements shows combining capacities of 1 and 2. It has more than one combining capacity.

An element which shows more than one combining capacity is said to have a variable valence or variable combining capacity.

Examine the following compounds:

$FeCl_2$	and	$FeCl_3$
FeO	and	Fe_2O_3
Hg_2Cl_2	and	$HgCl_2$
Hg_2O	and	HgO

Iron and mercury also have variable valencies. What are the combining capacities shown by each of these elements?

8.2 Combining Capacities of Some Elements

In Tables 8.2 and 8.3 the common combining capacities of some elements are given. You will notice that some atoms have only one combining capacity and some have more than one.

Table 8.2 Combining capacities of atoms of metallic elements

Element		*Combining capacity*
Potassium	decreasing electropositive nature	1
Sodium		1
Barium		2
Calcium		2
Magnesium		2
Aluminium		3
Manganese		2, 4, 7
Zinc		2
Iron		2, 3
Tin		2, 4
Lead		2, 4
Copper		1, 2
Mercury		1, 2
Silver		1

Table 8.3 Combining capacities of atoms of non-metallic elements

Element		*Combining capacity*
Hydrogen	decreasing electropositive nature; increasing electronegative nature	1
Phosphorus		3, 5
Carbon		2, 4
Sulphur		2, 4, 6
Iodine		1
Bromine		1
Chlorine		1
Nitrogen		1, 2, 3, 4, 5
Oxygen		2

Oxidation numbers

We have learned to express the combining capacity or valence of an element with reference to hydrogen which has been assigned a combining capacity of 1. The valence or combining capacity so expressed, does not convey any idea as to whether the atoms of an element form the electropositive part or the electronegative part of the molecule. Chemists, therefore, devised a term which will describe the combining capacity of the element and also indicate the positive and negative nature of its atoms in the compounds. This term is called the oxidation number.

The oxidation number is related to, but not identical with valence or combining capacity.

The use of oxidation numbers

In the compounds formed by the combination of metals with non-metals, the metals always show electropositive character. The combining capacities of the metals in such cases are expressed by using positive oxidation numbers. The combining capacities of the non-metallic elements are expressed by using negative oxidation numbers.

For example, in NaCl, the oxidation number of sodium is + 1, and that of chlorine is - 1. In CaO, the oxidation number of calcium is + 2 and the oxidation number of oxygen is - 2.

In compounds formed by the combination of one non-metal with another, the more electropositive element is assigned a positive oxidation number and the other is assigned a negative oxidation number.

For example, in HCl, hydrogen is assigned an oxidation number of + 1 and chlorine is assigned an oxidation number of - 1.

The use of the terms combining capacity and oxidation number may be illustrated by an example: Iron has a combining capacity of 2 and an oxidation **number of** + 2 in $FeCl_2$; it has a combining capacity of 3 and an oxidation number of **+ 3 in $FeCl_3$.**

Oxidation numbers of some elements

In Tables 8.4 and 8.5 the common oxidation numbers of some elements are given. You will notice that some atoms have only one oxidation number and some have more than one.

Table 8.4 Oxidation numbers of some metallic elements

Element	*Symbol with oxidation number*
Potassium	K^{1+}
Sodium	Na^{1+}
Calcium	Ca^{2+}
Barium	Ba^{2+}
Magnesium	Mg^{2+}
Aluminium	Al^{3+}
Manganese	$Mn^{2+, 4+, 7+}$
Zinc	Zn^{2+}
Iron	$Fe^{2+, 3+}$
Tin	$Sn^{2+, 4+}$
Lead	$Pb^{2+, 4+}$
Copper	$Cu^{1+, 2+}$
Mercury	$Hg^{1+, 2+}$
Silver	Ag^{1+}

Metallic elements are always assigned positive oxidation numbers.

Table 8.5 Oxidation numbers of some non-metallic elements

Element	*Symbol with oxidation number*
Hydrogen	H^{1+}
Phosphorus	$P^{3+, 5+}$
Carbon	$C^{2+, 4+}$
Sulphur	$S^{2-, 4+, 6+}$
Iodine	I^{1-}
Bromine	Br^{1-}
Chlorine	Cl^{1-}
Nitrogen	$N^{3-, 1+, 2+, 3+, 4+, 5+}$
Oxygen	O^{2-}

8.3 Writing Formulae of Compounds

For convenience we shall discuss the formula writing and the naming of compounds in the following order:

Binary compounds
Acids and acid radicals
Salts
Hydroxides

Formulae of binary compounds

Binary compound: A binary compound is a compound which contains two elements only.

Metallic binary compound: Binary compounds containing a metal are called metallic binary compounds.

Non-metallic binary compound: Binary compounds containing non-metallic elements only are called non-metallic binary compounds.

Order of symbols in a formula: Usually the symbol of the more, electropositive element present in the compound with positive oxidation number is written in front of the symbol of the element with negative oxidation number. (One exception is NH_3 in which the symbol of nitrogen with negative oxidation number (N) is written first.)

First symbol	Second symbol
positive oxidation number	negative oxidation number

Writing formulae of binary compounds: In the formula of a compound, the algebraic sum of the oxidation numbers must be equal to zero.

The sum of positive oxidation numbers	+	The sum of negative oxidation numbers	= 0

By applying this rule we can find out the number of atoms of each element which should be present in a formula. Let us begin with an example.

Example : When H and S combine with each other what will be the possible formula?

The oxidation number of H is + 1

The oxidation number of S is - 2

$H^{1+} + H^{1+} + S^{2-} = (H_2S)^0$

$\therefore$ The formula $= H_2S$

Other examples are:

Na and Cl	$Na^{1+} + Cl^{1-}$	=	$(NaCl)^0$	$\therefore$ The formula is NaCl
Ca and O	$Ca^{2+} + O^{2-}$	=	$(CaO)^0$	$\therefore$ The formula is CaO
Mg and Cl	$Mg^{2+} + Cl^{1-} + Cl^{1-}$	=	$(MgCl_2)^0$	$\therefore$ The formula is $MgCl_2$
Al and O	$Al^{3+} + Al^{3+} + O^{2-} + O^{2-} + O^{2-}$	=	$(Al_2O_3)^0$	$\therefore$ The formula is Al_2O_3

Naming binary compounds

I. Binary compounds in which the first element has a fixed oxidation number

The elements such as H^{1+}, K^{1+}, Na^{1+}, Cl^{1-}, Ca^{2+}, Zn^{2+} etc., have fixed oxidation numbers, Binary compounds in which the more electropositive element with fixed oxidation number is in the first place of the formula are named thus :

Name of the first element	Name of the second element ending in - **ide**

(Usually the change to - **ide** is in the second syllable of the name.)

Example:

H_2S	Hydrogen sulphide
CaO	Calcium oxide
$BaCl_2$	Barium chloride
AlN	Aluminium nitride

II. Binary compounds in which the first elements has a variable oxidation numbers

The elements such as C, S, Fe, Cu have variable oxidation numbers.

When the more electropositive element in the compound has variable oxidation numbers the name should be given thus:

1. For the naming of non-metallic binary compounds, Greek prefixes (e.g., mono-, di-, tri-, etc.) are used to indicate the number of atoms of each element in the compound. The name of the second element is ended with the syllable **-ide.**

   (1 = mono, 2 = di, 3 = tri , 4 = tetra, 5 = penta)

In those cases where there is only one atom of the first element, the use of the prefix mono is not necessary. The absence of a prefix to the name of the first element indicates that only one atom of the first element is present

di, tri, tetra, etc.	mono, di, tri, tetra, penta, etc,
Name of the first element	Name of the second element ending in **-ide**

Example:

N_2O	dinitrogen monoxide (dinitrogen oxide)
NO	nitrogen monoxide (nitrogen oxide)
N_2O_3	dinitrogen trioxide.
NO_2	nitrogen dioxide
N_2O_5	dinitrogen pentoxide
CCl_4	carbon tetrachloride

2. For the metallic binary compound, the name of the more electropositive metallic element with variable oxidation number is given first, followed by the Roman Numeral to state the oxidation number of the metallic element in the compound and the name of the second element ending **-ide**, is added.

Oxidation number of the first element in Roman Numeral

Name of the first element

Name of the second element ending in **-ide**

Example:

		Pronunciation
$FeCl_2$	iron (II) chloride	iron two chloride
$FeCl_3$	iron (III) chloride	iron three chloride
PbO	lead (II) oxide	lead two oxide
*Cu_2O	copper (I) oxide	copper one oxide
CuO	copper (II) oxide	copper two oxide
HgO	mercury (II) oxide	mercury two oxide

---

* Note that although two atoms of copper with oxidation number + 1 are present the prefix di-is not used for the metallic part of the name.

Formulae of some acids

The formulae of acids which you will come across in this course, are given below.

The names and formulae of some acids

Hydrogen chloride (g) Hydrochloric acid (aq)	HCl	These acids do not contain oxygen and their names start with "Hydro".
Hydrogen bromide (g) Hydrobromic acid (aq)	HBr	
Hydrogen iodide (g) Hydriodic acid (aq)	HI	
Hydrogen sulphide (g)	H_2S	

Nitrous acid	HNO_2	These acids contain oxygen and the word "Hydro" is not present in their names.
Nitric acid	HNO_3	
Sulphurous acid	H_2SO_3	
Sulphuric acid	H_2SO_4	
Carbonic acid	H_2CO_3	
Chloric acid	$HClO_3$	
Phosphoric acid	H_3PO_4	

Formulae of acid radicals

Derivation of acid radicals from the formulae of acids

HCl: The formula of hydrochloric acid is HCl. In HCl only one atom of hydrogen, H, is present. When H is removed from the formula, the remaining part is-Cl. This represents the acid radical of hydrogen chloride or hydrochloric acid.

$$HCl \longrightarrow -Cl \quad \text{(or)} \quad Cl^{1-}$$

H_2SO_4: The formula of sulphuric acid is H_2SO_4 in which there are two H atoms. When one H is removed from the formula, the remaining part is $-HSO_4$. This represents one acid radical of sulphuric acid. In $-HSO_4$ one H atom is still present. When this H atom is also removed, $-HSO_4$ becomes $>SO_4$. This represents another acid radical of sulphuric acid.

$$H_2SO_4 \begin{cases} \nearrow -HSO_4 & \text{(or)} \quad (HSO_4)^{1-} \\ \searrow >SO_4 & \text{(or)} \quad (SO_4)^{2-} \end{cases}$$

In acids like H_2SO_4 where there are two H atoms we get two acid radicals ($-HSO_4$ and $>SO_4$). In acids like HCl where only one H atom is present, we get only one acid radical (Cl^{1-}).

The formula of an acid radical is, therefore, that part of the formula of the acid that remains after removing one or more hydrogen atoms.

If only one H atom is present in the formula of the acid, only one formula is possible for the acid radical.

If two H atoms are present in the formula of the acid, two formulae corresponding to two radicals are possible, and so on.

Determination of the oxidation numbers of acid radicals

The oxidation number of the acid radical in an acid is a negative number that is numerically equal to the number of the H atoms that has to be removed from the formula of the corresponding acid to arrive at the formula of the acid radical.

Example: - Cl is obtained by taking away one H atom from HCl and the oxidation number of - Cl is - 1.

$$HCl \xrightarrow{(-H)} -Cl^{1-}$$

Example: $> SO_4$ is obtained by taking away two H atoms from H_2SO_4 and the oxidation number of $> SO_4$ is - 2.

$$H_2SO_4 \xrightarrow{(-H)} (HSO_4)^{1-}$$
$$H_2SO_4 \xrightarrow{(-2H)} (SO_4)^{2-}$$

Derivation of the name of the acid radical from that of the acid

The names of the acids, which have no oxygen atoms in their molecules begin with the word "Hydro" and end with the word "ic" e.g., HCl is called hydrochloric acid.

In naming the acid radicals of these acids the word "hydro" is omitted in the name of acid radical and -ic is changed to -ide, e.g., the acid radical of hydrochloric acid is chloride.

Hydro **-ic** acid	$\longrightarrow$	**-ide**
Acid		Acid Radical
Hydrochloric acid (HCl)	$\longrightarrow$	chloride (-Cl) Cl^-
Hydrobromic acid (HBr)	$\longrightarrow$	bromide (- Br) Br^-
Hydriodic acid (HI)	$\longrightarrow$	iodide (- I) I^-
Hydrogen sulphide (H_2S)	$\longrightarrow$	sulphide (>S) S^{2-}

The relation between the names of the acids which have oxygen atoms in their molecules and the names of the acid radicals is as follows :

Name of the acids		Name of the acid radicals
-ous acid	$\longrightarrow$	**-ite**
-ic acid	$\longrightarrow$	**-ate**
Example: Nitrous acid (HNO_2)	$\longrightarrow$	nitrite ($-NO_2$)
Nitric acid (HNO_3)	$\longrightarrow$	nitrate ($-NO_3$)
Sulphurous acid (H_2SO_3)	$\longrightarrow$	sulphite ($>SO_3$)
Sulphuric acid (H_2SO_4)	$\longrightarrow$	sulphate ($> SO_4$)

The acid radicals with H: When H is present in the formula of an acid radical, the word hydrogen is placed before the name of the acid radical.

Example : HSO_3^{1-} = hydrogensulphite

HSO_4^{1-} = hydrogensulphate

HCO_3^{1-} = hydrogencarbonate

From the above principles, you can derive the formula, the name and the oxidation number of an acid radical from the name and the formula of the corresponding acid.

The name and the oxidation number of the acid radical may be derived from the name of acid as shown in Table 8.6

Table 8.6 Acids and acid radicals

Name of acid	*Formula of acid*	*Acid radical*	*Name of acid radical*	*Oxidation number*	*Number of acid radical*
Hydrochloric acid	HCl	- Cl	chloride	- 1	1
Hydrobromic acid	HBr	- Br	bromide	- 1	1
Hydriodic acid	HI	- I	iodide	- 1	1
Nitrous acid	HNO_2	- NO_2	nitrite	- 1	1
Nitric acid	HNO_3	- NO_3	nitrate	- 1	1
Carbonic acid	H_2CO_3	- HCO_3	hydrogencarbonate	- 1	2
		> CO_3	carbonate	- 2	
Sulphurous acid	H_2SO_3	- HSO_3	hydrogensulphite	- 1	2
		> SO_3	sulphite	- 2	
Sulphuric acid	H_2SO_4	- HSO_4	hydrogensulphate	- 1	2
		> SO_4	sulphate	- 2	
Chloric acid	$HClO_3$	- ClO_3	chlorate	- 1	1
Phosphoric acid	H_3PO_4	- H_2PO_4	dihydrogenphosphate	- 1	3
		> HPO_4	hydrogenphosphate	- 2	
		→ PO_4	phosphate	- 3	

Ammonium radical and its oxidation number: The formula of ammonium radical is NH_4. Its oxidation number is + 1.

Formula and oxidation number ⟶ NH_4^{1+}

Formula of salts

Formula : The formula of a salt consists of two parts. The first part is the metal atom or the ammonium radical. The second part is the acid radical.

First part | Second part

Metal atom or NH_4^+	Acid radical

Writing the formulae of salts: When writing the formula of a salt, we should observe the following rule.

The positive oxidation number of the metal atoms or the ammonium radical	+	The negative oxidation number of the acid radicals	=	0

Examples for writing the formulae of salts

Let us first consider how to write the formula of sodium sulphate. To get the formula of sodium sulphate, we must combine the Na^{1+} with the SO_4^{2-}.

In order to make the sum of the total oxidation number equal to zero, we must combine $2Na^{1+}$ with SO_4^{2-}

Then the sum of the oxidation numbers = 2 (+ 1) + (- 2) = O

$\therefore$ The required formula $= 2Na^{1+} + SO_4^{2-} = Na_2SO_4$

Other examples are

(i) $K^{1+} + NO_3^{1-} = KNO_3$

(ii) $Ca^{2+} + SO_4^{2-} = CaSO_4$

(iii) $Ca^{2+} + HCO_3^{1-} + HCO_3^{1-} = Ca(HCO_3)_2$

Read as Ca-H-C-O-3 twice

(iv) $NH_4^{1+} + NH_4^{1+} + SO_4^{2-} = (NH_4)_2SO_4$

Read as N-H-4 twice S-O-4

Naming salts

Naming the salt containing a metal atom with constant oxidation number or an ammonium radical

The name of the salt begins with the name of the metal or ammonium radical, followed by the name of the acid radical as shown below.

The name of the salt =	The name of the metal atom of NH_4^+	The name of the acid radical

Example :

formula	Name
$Ca(NO_3)_2$	Calcium nitrate
KNO_3	Potassium nitrate
$(NH_4)_2SO_4$	Ammonium sulphate

Naming the salt in which the metal atom has a variable oxidation number

The name beings with the name of the metal, with Roman Numeral, indicating the oxidation number of the metal atom and followed by the name of the acid radical.

The name of the salt = $\boxed{\text{The name of the metal atom}}$ $\left\{\text{Roman Numeral}\right\}$ $\boxed{\text{The name of the acid radical}}$

Note The Roman Numeral represents the oxidation number of the metal atom.

Example:

$FeSO_4$	Iron (II) sulphate
$Fe_2(SO_4)_3$	Iron (III) sulphate

Formula of hydroxide

All hydroxides include one or more - OH radicals of oxidation number -1 in their formulae.

The general formula can be diagrammatically represented as follows :

Metal atom or NH_4^+	OH radical

Writing the formula of hydroxide

The positive oxidation number of the metal atom or the ammonium radical	+	Total negative oxidation number of the OH radicals	= 0

Naming hydroxides: Hydroxides are named in such the same way as the naming of salts, but the name of acid radical is replaced by the word hydroxide, e.g. :

The metal atom with fixed oxidation number	$NaOH$	sodium hydroxide
	$Ca(OH)_2$	calcium hydroxide
The ammonium radical	NH_4OH	ammonium hydroxide
The metal atom with variable oxidation number	$Fe(OH)_2$	iron (II) hydroxide
	$Fe(OH)_3$	iron (III) hydroxide

The common names of some compounds : What you have seen above are the chemical names of some typical compounds. You should note that chemical compounds have old names and common names used in every day life as well as in industries and chemical literature (Table 8.7 and 8.8).

Table 8.7 Common names of some substances

Formula	*Chemical name*	*Common name*
KOH	potassium hydroxide	caustic potash
NaOH	sodium hydroxide	caustic soda
NaCl	sodium chloride	common salt
CaO	calcium oxide	quick lime
$Ca(OH)_2$	calcium hydroxide	slaked lime
$CaCO_3$	calcium carbonate	limestone, marble
Na_2CO_3	sodium carbonate	washing soda

Table 8.8 Old names of some substances

Formula	*New name*	*Old name*
$FeCl_2$	iron (II) chloride	ferrous chloride
$FeCl_3$	iron (III) chloride	ferric chloride
PbO	lead (II) oxide	lead monoxide
PbO_2	lead (IV) oxide	lead peroxide (or) lead dioxide
$CuSO_4$	copper (II) sulphate	cupric sulphate (or) copper sulphate
HgO	mercury (II) oxide	mercuric oxide
MnO_2	manganese (IV) oxide	manganese dioxide
N_2O	dinitrogen oxide	nitrous oxide
NO	nitrogen monoxide	nitric oxide
NO_2	nitrogen dioxide	nitrogen peroxide

8.4 Balancing Equations

The systematic method of balancing equations will be discussed in the higher class. At the present stage your teacher will show you some helpful ways of balancing chemical equations. The skill in balancing equations can be acquired only by practice.

8.5 Writing Specific Equations from General Reactions

Chemists have determined the products of the different reactions which substances undergo by actual experiments. By comparing the reactions, they have found that there are some regular patterns in the way products are formed when certain substances react together or when certain substances, such as carbonates, hydroxides, decompose.

Example 1

All metal hydroxides, except sodium hydroxide and potassium hydroxide, decompose on heating in the following regular pattern:

metal hydroxide	$\longrightarrow$	**metal oxide**	+	**water**
calcium hydroxide	$\longrightarrow$	calcium oxide	+	water
$Ca(OH)_2$	$\longrightarrow$	CaO	+	H_2O
copper (II) hydroxide	$\longrightarrow$	copper (II) oxide	+	Water
$Cu(OH)_2$	$\longrightarrow$	CuO	+	H_2O

Example 2

All metal carbonates, except sodium carbonate and potassium carbonate, decompose on heating in the following regular pattern:

metal carbonate	$\longrightarrow$	**metal oxide**	+	**carbon dioxide**
copper (II) carbonate	$\longrightarrow$	copper (II) oxide	+	carbon dioxide
$CuCO_3$	$\longrightarrow$	CuO	+	CO_2
calcium carbonate (limestone)	$\longrightarrow$	calcium oxide (quicklime)	+	carbon dioxide
$CaCO_3$	$\longrightarrow$	CaO	+	CO_2

Example 3

All carbonates react with dilute acids forming salts, water and carbon dioxide.

As an example we shall consider the reaction between calcium carbonate and hydrochloric acid. According to the general reaction, the equation will be

calcium carbonate + hydrochloric acid $\longrightarrow$ salt + water + carbon dioxide

Now let us consider what salt should be formed in this reaction.

(1) The metallic atom, present in the reactant, is calcium. Therefore the salt in the product will be a Ca salt.

$$CaCO_3 + HCl \longrightarrow \text{Ca salt} + H_2O + CO_2$$

(2) To form a salt, the Ca radical must combine with an acid radical. Since we are using hydrochloric acid, the acid radical must be - Cl.

(3) So the salt is calcium chloride. Its formula is $CaCl_2$, since the oxidation number of Ca is + 2 and that of Cl is - 1.

So, the equation for the above reaction can be written as below.

calcium carbonate	+	hydrochloric acid	$\longrightarrow$	calcium chloride	+	water	+	carbon dioxide
$CaCO_3$	+	$2HCl$	$\longrightarrow$	$CaCl_2$	+	H_2O	+	CO_2

In this way you can write an equation for any carbonate reacting with any dilute acid.

To get practice and improve your skill in writing specific equations from the general reactions, try to write by yourself the equations for the following reactions.

(1) Sodium carbonate and dilute hydrochloric acid.
(2) Magnesium carbonate and dilute sulphuric acid.
(3) Copper (II) carbonate and dilute sulphuric acid.

SUMMARY

In this chapter, chemical reactions can be described by means of chemical equations using symbols and formulae.The concepts of combining capacity of elements in compounds have been introduced. The combining capacity or valence of an element is represented by the number of atoms of hydrogen, chlorine or sodium that combine with one atom of that element. Fixed combining capacity of certain elements such as calcium and magnesium always show a combining capacity of 2 in their compounds can be given. Variable combining capacities of certain elements such as copper which forms two oxides: copper (I) Oxide, Cu_2O and copper (II) oxide, CuO indicates copper has a combining capacity of 1 and 2. An element which shows more than one combining capacity is said to have a variable valence or variable combining capacity. The use of oxidation numbers in compounds formed by the combination of metals with non-metals where the metals always show electropositive character or positive oxidation number and non-metals show negative oxidation number.For example, in NaCl , the oxidation number of sodium is + 1 and that of chlorine is – 1.Next , formulae writing of binary compounds, acid and acid radicals, salts and hydroxides have been presented.Writing specific equations from general reactions have been depicted with appropriate examples.

Questions and problems.

1. Metallic oxides react with dilute acids in the following pattern:

$$\text{metallic oxide} + \text{acid} \longrightarrow \text{salt} + \text{water}$$

Now write equations in words and symbols for the reactions between each oxide in column I and each acid in column II.

Column I	Column II
sodium oxide	hydrochloric acid
calcium oxide	nitric acid
zinc oxide	sulphuric acid
copper (II) oxide	

2. Hydroxides react with dilute acids in the following pattern :

metallic hydroxide + acid $\longrightarrow$ salt + water

Write equations in words and symbols for the reactions between each hydroxide in column I and each acid in column II

Column I	Column II
potassium hydroxide	hydrochloric acid
calcium hydroxide	nitric acid
aluminium hydroxide	sulphuric acid
ammonium hydroxide	

3. Metals react with hydrochloric acid or dilute sulphuric acid in the following pattern :

metal + acid $\longrightarrow$ salt + hydrogen

Write equations in words and symbols for the reactions between each metal in column I and each acid in column II.

Column I	Column II
magnesium	dilute hydrochloric acid
aluminium	dilute sulphuric acid
zinc	dilute hydrobromic acid

4. Metallic sulphides react with dilute acids in the following pattern:

metallic sulphide + acid $\longrightarrow$ salt + hydrogen sulphide

Write equations in words and symbols for the reactions between each sulphide in column I and each acid in column II

Column I	Column (II)
lead (II) sulphide	dilute hydrochloric acid
iron (II) sulphide	dilute sulphuric acid
zinc sulphide	

5. Metallic sulphites react with dilute acids to produce salts, water and sulphur dioxide. Write the equations for sodium, calcium and aluminium sulphites, when each reacts with each of the following acids. Hydrochloric aicd, sulphuric acid, nitric acid.

6. Metals react with concentrated sulphuric acid to form salts, water and sulphur dioxide. Write equations for the reactions between
   (a) zinc and concentrated sulphuric acid.
   (b) magnesium and concentrated sulphuric acid.
   (c) copper and concentrated sulphuric acid.
7. Complete the following equations.
   (a) $Na_2O + HCl \longrightarrow ?$
   (b) $K_2O + H_2SO_4 \longrightarrow ?$
   (c) $CaO + HNO_3 \longrightarrow ?$
   (d) $Al_2O_3 + H_2SO_4 \longrightarrow ?$
   (e) $CaO + H_3PO_4 \longrightarrow ?$
   (f) $Al_2O_3 + H_3PO_4 \longrightarrow ?$
   (g) barium hydroxide + dilute sulphuric acid $\longrightarrow$?
   (h) zinc hydroxide + dilute hydrochloric acid $\longrightarrow$?
   (i) iron (III) hydroxide + dilute nitric acid $\longrightarrow$?
   (j) copper (II) hydroxide + dilute sulphuric acid $\longrightarrow$?
   (k) $Mg + HCl \longrightarrow ?$
   (l) $Al + H_2SO_4$ (dilute) $\longrightarrow$?
   (m) $Cu + H_2SO_4$ (concentrated) $\longrightarrow$?
   (n) iron (II) sulphide + sulphuric acid (dilute) $\longrightarrow$?
   (o) $Na_2SO_3 + HCl \longrightarrow ?$
8. Write TRUE or FALSE for each of the following statements.
   (a) The combining capacity or valency of an element is represented by the number of atoms of hydrogen, chlorine or sodium that combine with one atom of that element.
   (b) Combining capacities of atoms of metallic element cannot be more than one.
   (c) Combining capacities of atom of non-metallic element cannot be more than one.
   (d) The oxidation number of metallic elements is not related to combining capacity of the elements.
   (e) In NaCl, the oxidation number of sodium is + 1 and that of chlorine - 1.
   (f) The oxidation number of Al is 3.
   (g) In the formula of a compound, the sum of the oxidation number must be equal to zero.
   (h) Metallic oxide reacts with dilute acid to produce salt and hydrogen.

(i) Metallic elements are always assigned positive oxidation numbers in their compounds.

(j) Metal free compound is sodium chloride.

(k) A molecule of ammonia contains an atom of nitrogen and three atoms of oxygen.

9. Fill in the blanks with a suitable word or words or symbols or number.

(a) In H_2O, two atoms of .... combine with one atom of ....

(b) Aluminium, in aluminium chloride, $AlCl_3$, shows a combining capacity of ....

(c) In NaCl, the oxidation number of sodium is .... and that of chlorine is ....

(d) A.......... compound is a compound which contains two elements only.

(e) Binary compounds containing a metal are called .... binary compounds.

(f) Oxygen and sulphur have .... oxidation numbers and sodium and potassium have .... oxidation numbers.

(g) All carbonates react with dilute acids forming salt, water and ....

(h) An element which shows more than one combining capacity is said to have a .... valence.

(i) The formula of sodium oxide is ....

(j) The name of $FeCl_2$ is ....

10. Select the correct word or words given in the brackets.

(a) In H_2S and Na_2S, sulphur shows a combining capacity of (1,2,3).

(b) An element which shows more than one combining capacity is said to have a (fixed, constant, variable) valence.

(c) The common name for sodium hydroxide is (caustic potash, caustic soda, common salt).

(d) The old name of iron (II) chloride is (ferrous chloride, ferric chloride, lead chloride),

(e) The common name for calcium hydroxide is (quicklime, slaked lime, limestone).

11. Match each of the items given in list A with the appropriate item given in list B.

List A	List B
(a) Calcium carbonate reacts with hydrochloric acid	(i) washing soda
(b) The common name for sodium chloride	(ii) +1 and -2
(c) The common name for sodium carbonate	(iii) manganese dioxide
(d) The oxidation number of hydrogen and sulphur in H_2S	(iv) common salt
(e) The old name for manganese (IV) oxide	(v) calcium chloride, carbon dioxide and water are formed

12. Answer the following questions.

(a) For each of the following word equations, write a balanced symbol equation.

(i) zinc oxide + nitric acid ⟶ zinc nitrate + water

(ii) aluminium + sulphuric acid(dil) ⟶ aluminium sulphate + hydrogen

(b) Balance the following equations in **words** and **symbols.**

(i) $CaCO_3 + HCl \longrightarrow CaCl_2 + H_2O + CO_2$

(ii) $Ca(OH)_2 \longrightarrow CaO + H_2O$

(c) Complete the following word and symbol equations.

(i) sodium oxide + ... ⟶ sodium chloride + ...

(ii) .... + H_2SO_4 (conc) ⟶ $CaSO_4$ + ...

(d) Complete the following table, giving the name or formula of each compound as appropriate.

Sr. No	Chemical name	Chemical formula
1	Barium hydroxide	
2		$CaCO_3$
3	Iron (II) sulphate	

(e) Write equations in words and symbols for the reactions between

(i) zinc and concentrated sulphuric acid.

(ii) sodium carbonate and dilute hydrochloric acid.

(f) Write down the formula of each of the following compounds.

(i) common salt

(ii) quick lime

(iii) caustic soda

(iv) caustic potash

13. Complete the following equations.

(a) $Cu(OH)_2 \longrightarrow ?$

(b) $CaCO_3 + HCl \longrightarrow ?$

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CHAPTER 9

THE MOLE CONCEPT

In this chapter we shall describe the chemist's ideas of measuring the amount of substances. Then, based upon these ideas, Avogadro's Number will be introduced, to show the chemist's way of counting atoms and molecules. We shall also describe the unit amount of substance - the mole or gram molecule which represents both the number of particles and the mass in grams in calculations based on chemical equations.

9.1 Measurement of the Amount of Substance

In every day life and in your school laboratory, you will notice that the amount of some substances are measured by mass, some by volume and some by number, depending upon the more convenient way to measure.

When carbon is burnt in oxygen, carbon dioxide is formed. The chemical equation for this reaction is

$$C(s) \quad + \quad O_2(g) \quad = \quad CO_2(g)$$

This equation shows that each atom of carbon combines with one molecule of oxygen forming one molecule of carbon dioxide. One molecule of oxygen consists of two atoms of oxygen and one molecule of carbon dioxide consists of one atom of carbon and two atoms of oxygen.

Chemists usually measure the amount of atoms and molecules in terms of numbers.

9.2 The Amount of Substances and the Number of Atoms or Molecules

Let us study the numbers and the masses of atoms. (Relative masses are taken as the nearest whole number.)

The relative mass of one atom of H is 1.

The relative mass of one atom of C is 12.

The relative mass of one atom of O is 16.

The relative mass of one atom of Mg is 24.

The relative mass of one molecule of H_2 is 2.

The atoms and the molecules in the above examples are not equal in mass, but they are equal in numbers, because there is one atom or one molecule in each example.

Therefore, to compare the amounts of atoms and molecules it is more convenient to compare their number rather than their mass. It is inconvenient to compare the atoms and the molecules in terms of mass, because the atoms of different elements have different masses.

9.3 The Amount of Substance from the Chemist's Point of View

We have learned that all substances are composed of atoms or molecules. When chemists see a pile of any substance, they see, in their mind's eye, that the pile of the substance is nothing but a collection of atoms or molecules.

For chemists, therefore:

(1) The amount of a substance is related to the number of particles (atoms, molecules, etc.) present in the substance.

(2) One way of comparing the amounts of different substances is to compare the number of particles present in each of the substances.

9.4 Measurement of the Amount of Substance in terms of Number of Atoms

Since atoms and molecules are extremely small, we cannot count them one by one to compare their numbers. It is possible, however; to compare different sets of equal numbers of particles. Chemist's choice for the equal number of particles in a set is as follows:

The relative atomic mass of hydrogen is 1.

It is assumed that the number of atoms in 1 g of hydrogen is L.

$\therefore$ The Mass of L atoms of H is 1 g.

The relative atomic mass of carbon is 12.

$\therefore$ One C atom is 12 times heavier than ' atom of hydrogen.

$\therefore$ The Mass of L atoms of C is 12 g.

The relative atomic mass of oxygen is 16.

$\therefore$ One O atom is 16 times heavier than 1 atom of hydrogen.

$\therefore$ The Mass of L atoms of O is 16 g.

The relative atomic mass of magnesium is 24.

$\therefore$ One Mg atom is 24 times heavier than 1 atom of hydrogen.

$\therefore$ The Mass of L atoms of Mg is 24g.

But 1, 12, 16 and 24 are the relative atomic masses of H, C, O and Mg respectively.

Therefore, 1 g of H, 12 g of C, 16 g of O and 24 g of Mg are the relative atomic masses of the respective element expressed in grams.

(a) What is the conclusion you can draw from the above facts?

If the relevant conclusion still eludes you, try to answer the following questions (b) and (c) then try to answer the question (a) again.

(b) What is the number of atoms present in the relative atomic mass in grams of each element?

(c) What is the relation between the numbers of atoms present in the relative atomic mass in grams of each element?

Since the number of atoms present in the relative atomic mass in grams of each element is L, it can be concluded that :

The number of atoms present in the relative atomic mass in grams is the same for all elements.

9.5 Measurement of the Amount of Substance in terms of Number of Molecules

The relative atomic mass of hydrogen is 1.

It is assumed that the number of atoms in 1 g of hydrogen is L.

$\therefore$ The Mass of L atoms of H is 1 g.

One molecule of hydrogen consists of 2 atoms of hydrogen.

$\therefore$ The relative molecular mass of hydrogen is 2.

$\therefore$ One molecule of H_2 is 2 times heavier than 1 atom of H.

$\therefore$ The Mass of L molecules of H_2 is 2 g.

The relative molecular mass of H_2O is 18.

$\therefore$ One molecule of H_2O is 18 times heavier than 1 atom of H.

$\therefore$ The Mass of L molecules of H_2O is 18 g.

In the same way it can be shown that there will be L molecules in the .relative molecular mass in grams of every molecular substance.

Therefore, we can say in general that:

> The number of molecules present in the relative molecular mass in grams is the same for all elements and compounds, i.e., the molecular mass in grams of all elements and compounds contain the same number of molecules.

9.6 The Avogadro's Number or the Avogadro's Constant (L)*

As described above, the number of atoms present in the relative atomic mass in grams is the same for all elements. By experimental investigation it was found that the numerical value of this number is 6.02×10^{23}. This number is known as the **Avogadro's Number** and is represented by the symbol L.

$$L = 6.02 \times 10^{23}$$

Since the Avogadro's Number is constant for all elements, it is also called the **Avogadro's Constant** and may be defined by referring to the atomic mass in grams of any element. However, since ^{12}C has been chosen as the reference element, Avogadro's Number or Constant is defined by referring to ^{12}C as follows:

> The Avogadro's Number or Constant is the number of carbon atoms present in exactly 12 g of ^{12}C.

Chemists have chosen the Avogadro's Number as the standard for measuring the amount of atoms and molecules.

Remember that the Avogadro's Number is related to the relative atomic mass in grams or the relative molecular mass in grams.

The relative atomic mass in grams of any element contains 6.02×10^{23} atoms, the number being equal to the Avogadro's Number. The relative molecular mass in grams of any element or compound contains 6.02×10^{23} molecule, the number being equal to the Avogadro's Number.

We shall see below that we can count the number of atoms and molecules by weighing out multiples or fractions of the relative atomic masses or the relative molecular masses.

9.7 Counting the Number of Atoms and Molecules by Weighing

The exact relative atomic mass in grams of Mg is 24.3 g.

$\therefore$ 24.3 g of magnesium contains 6.02×10^{23} magnesium atoms.

* The numerical value of the Avogadro's Number was first determined by Loschmidt and is now represented by the letter L to honour him.

In order to get 6.02×10^{23} atoms of magnesium, we can just weigh out exactly 24.3 g of magnesium using a balance.

Or, if we want $\frac{6.02 \times 10^{23}}{2}$ atoms of magnesium, we can just weigh out

$\frac{24.3}{2}$ g = 12.15 g of magnesium using a balance.

In the same way, if we want $2 \times 6.02 \times 10^{23}$ atoms of magnesium, we can weigh out 2x 24.3 g of magnesium using a balance.

This is the chemist's way of counting the required numbers of atoms by the weighing the fractions or the multiples of the relative atomic masses in grams. The chemist can count the number of molecules by weighing in the same way.

Now, there are 6.02×10^{23} molecules in 32 g or the relative molecular mass in grams of oxygen.

Hence, to obtain 6.02×10^{23} molecules of oxygen, we must weigh out 32 g of oxygen.

Or, if we want $\frac{6.02 \times 10^{23}}{2}$ molecules of oxygen, we must weigh out $\frac{32}{2}$ g = 16g of oxygen.

From the above illustrations, we can see that to obtain the required numbers of atoms or molecules of substances, we have to weigh out multiples or fractions of the relative atomic or molecular masses in grams of the substances which contain the required numbers of atoms or molecules.

9.8 The Unit Amount of a Substance - The Mole

For chemists, the amount of a substance means the number of particles present in the substance. The standard measure is the Avogadro's Number which is the number of atoms in the relative atomic mass in grams of ^{12}C. The unit amount of a substance is, therefore, the amount of the substance which contains the Avogadro's Number of its particles. This unit amount of a substance is called the mole.

Thus, the mole is defined as follows :

One mole of a substance is the amount of that substance which contains the same number of particles (atoms, molecules, etc.) as there are atoms in 12 g of ^{12}C.

The difference between one mole of atoms and one mole of molecules

Let us take oxygen as an example. One mole of oxygen atoms (O) contains the Avogadro's Number of the oxygen atoms. Since the relative atomic mass of O atom is 16, one mole of O atoms, that is the Avogadro's Number of O atoms, weighs 16 g.

Similarly, one mole of oxygen molecules contains the Avogadro's Number of the oxygen molecules. However, since two atoms of oxygen combine to form one oxygen molecule, the relative molecular mass of oxygen is $2 \times 16 = 32$. Therefore, one mole of oxygen molecules (O_2), that is the Avogadro's Number of oxygen molecules, weighs 32 g.

As shown above, one mole of atoms and one mole of molecules of the same element contain equal numbers of particles. But one mole of atoms and one mole of molecules are different in mass.

Thus, it is required to specify clearly what kind of particle is being considered in the calculation of one mole of a substance such as :

One mole of oxygen atoms (O)

(or)

One mole of oxygen molecules (O_2)

9.9 Correlation of the Number of Particles with their Masses by Mole Unit

The mole represent the Avogadro's Number of particles (atoms or molecules) as well as the total mass of these particles which is equal to relative atomic mass in grams or relative molecular mass in grams of the substance.

In this way the mole correlates the reacting number of atoms and molecules in a chemical equation with the actual masses in grams of the reacting substances. The actual masses in grams of the reacting substance must be weighed in a balance.

By using the mole units we can predict the masses of the reactants and the products that take part in a chemical reaction which is expressed by a chemical equation.

This can be illustrated by the following example.

Example:

When magnesium and sulphur are heated together, they react to form magnesium sulphide. What weight of sulphur is required to react with exactly 6 g of magnesium to form magnesium sulphide?

The chemical equation for this reaction is

$$Mg\,(s) \quad + \quad S\,(s) \quad = \quad MgS\,(s)$$

From the above chemical equation we can see that each atom of Mg reacts with 1 atom of S to form MgS.

In what mole ratio do Mg atoms react with S atoms?

The above chemical equation points out that Mg atoms and S atoms react in equal numbers to form MgS.

Therefore, L number of Mg atoms will react with L number of S atoms to form MgS.

Since L number of atoms is 1 mole of atoms, 1 mole of Mg atoms will react with 1 mole of S atoms to form MgS.

How many atoms are there in 6 g of Mg?

Since the relative atomic mass of Mg is 24, 1 mole (L numbers) of Mg atoms weighs 24 g.

$\therefore$ 24 g of Mg contains 1 mole (L number) of Mg atoms

$\therefore$ 6 g .... ?

$$= \frac{6}{24} = 0.25 \text{ mol of Mg atoms}$$

How many moles of S atoms react with 0.25 mole of Mg?

As we have already mentioned above, 1 mole of Mg atoms reacts with 1 mole of S atoms.

Therefore, 0.25 mole of Mg atoms will react with 0.25 mole of S atoms.

What is the mass of 0.25 mole of S atoms?

Since the relative atomic mass of S is 32, 1 mole (L numbers) of S atoms weighs 32 g.

Therefore, 0.25 mole (0.25 L numbers) of S atoms weighs 0.25 x 32 g = 8 g.

In this way we can predict that 8 g of S is required to react with 6 g of Mg to form MgS.

This is the reasoning followed in calculating the masses of the reactants and the products in a chemical equation. The main idea is :

(1) To find mole ratio of the reacting substances from the corresponding chemical equation.

(2) To find the weight ratio of the reacting substances from their mole ratio.

Reacting masses of substances in chemical equations

From the above examples, it is clear that each symbol or formula of a substance represents one mole of the substance.

Thus, a symbol represents one mole of atoms and a formula represents one mole of molecules in a molecular substance. In a non-molecular substance a formula represents one mole of particles as represented by the formula.

Formula = (a) 1 mole of molecules in a molecular substance

(b) 1 mole of particles represented by the formula in a non-molecular substance.

9.10 Calculation from Chemical Equations

In actual practice not all the steps shown above are required. The calculation is illustrated below.

Example: To find the weight of product from the weight of a reactant

When a piece of magnesium is burnt in air, 8 g of oxygen is used up. How many grams of magnesium oxide will be formed? (Mg = 24, O = 16)

	Calculation
Step 1	
Convert 8 g of oxygen into mole.	1 mol of O_2 = 32 g
The relative atomic mass of O = 16	Since 32 g of O_2 = 1 mol
∴ The relative molecular mass of O_2 is 32.	∴ 8 g of O_2 = $\frac{8}{32}$ = 0.25 mol of O_2
Step 2	
Write down the balanced equation and the mole ratio of the reactants and the products.	$2Mg\,(s) + O_2\,(g) = 2MgO\,(s)$ 0.25 mol = ?
Step 3	
Calculate the mole of MgO formed from 0.25 mole of O_2	From the equation, 2 moles of Mg and 1 mole of O_2 give rise to 2 moles of MgO.

∴ 0.25 mole of O_2 ... ... 2 x 0.25

= 0.5 mol of MgO

Step 4

Convert 0.5 mole of MgO into mass in grams.

Mg = 24 and O = 16

The relative formula mass of

MgO = 24 + 16 = 40

∴ 1 mol of MgO = 40 g

∴ 0.5 mol of MgO = 0.5 x 40 g

= 20 g

∴ 20 g of MgO is formed from 8 g of O_2

9.11 The Volume of One Mole of a Gas - The Molar Volume

In this section we shall describe the volume occupied by 1 mole of a gas.

Measurement of the amount of gases

Gases are comparatively light and difficult to handle so it is inconvenient to weigh a gas on a balance. Therefore they are usually measured by volume.

The volume, temperature and pressure of a gas

The volume of a given mass of gas changes with the change in temperature or pressure of the gas. It is therefore, important to specify the temperature and the pressure at which the volume of the gas is measured.

The density of a gas

The density of a gas usually expressed in grams per dm^3.The density of a gas is obtained, therefore, by dividing the mass by the volume.

Mathematical expression is:

$$\text{The density of a gas} = \frac{\text{The mass of the gas in grams}}{\text{The volume of the gas in dm}^3}$$

Since the density of a gas depends upon the volume which varies with temperature and pressure, the density of a gas also varies with temperature and pressure.

For easy reference and for convenience the densities of gases are measured under certain standard condition of temperature and pressure.

The standard temperature is 0 °C

The standard pressure is 760 mm Hg

They are referred to as STP meaning Standard Temperature and Pressure.

The volume of one mole of a gas at STP

The volume of one mole of a gas at STP in dm^3, can be calculated by dividing the relative mass of one mole of the gas by its density.

$$\text{The volume of one mole of a gas at STP} = \frac{\text{Relative mass of one mole of a gas}}{\text{The density of the gas}}$$

One mole of the gas can be calculated from the formula of the gas and its relative atomic mass.

The density of the gas is determined by experiment.

As an example, let us determine the volume at STP occupied by one mole of hydrogen.

The formula of hydrogen	$= H_2$
The relative atomic mass of H	$= 1.008$
The relative molecular mass of H_2	$= 2 \times 1.008 = 2.016$
$\therefore$ One mole of H_2	$= 2.016$ g
The density of hydrogen at STP	$= 0.098$ g dm^{-3} (found by experiment)

$$\therefore \text{The volume of one mole of } H_2 \text{ at STP} = \frac{\text{Relative mass of one mole of } H_2}{\text{The density of the gas}}$$

$$= \frac{2.016\ g}{0.098\ g\ dm^{-3}} = 22.4\ dm^3$$

$\therefore$ The volume of one mole of H_2 at STP $= 22.4\ dm^3$

The volume at STP of one mole of all gases

By finding out the volume at STP of one mole of oxygen, nitrogen, chlorine, carbon **dioxide**, etc., as you have done in your experimental chemistry book, it is found that the volume occupied by one mole of any gas is equal to 22.4 dm^3 at STP. This volume is known as the molar volume of the gas.

> One mole of any gas occupies a volume of 22.4 dm^3 at STP and this volume is called the molar volume of the gas.

In Section 9.5, we have learned that the molecular mass in grams (or mole) of all elements and compounds contain the same number of molecules. Since the mole of all gaseous elements or compounds occupy the same volume (22.4 dm^3) at STP, it follows that at STP one mole of gases occupies the same volume and contain the same number of molecules.

This fact is extended and summarized in **Avogadro's Theory or Law** which is stated as follows:

> At the same temperature and pressure equal volumes of all gases contain the same number of molecules.

The correlation of mass and volume by molar volume

By using the data that one mole of gas occupies a volume of 22.4 dm^3 at STP, we can convert the mass in grams of a gas to volume at STP and the volume of a gas at STP to mass in grams.

Example: **To find the volume of 8 g of oxygen at STP**

Calculation

The formula of oxygen gas = O_2

The relative atomic mass of O = 16

The relative molecular mass of O_2 = 2 x 16 = 32

$\therefore$ One mole of O_2 32 g

Mole of oxygen in 8 g of oxygen = $\frac{8g}{32g}$ = 0.25 mol

Since, 1 mol of O_2 occupies 22.4 dm^3 at STP

$\therefore$ 0.25 mol of O_2 .... 22.4 x 0.25 = 5.6 dm^3 at STP

Example: **To find the mass in grams of 2.8 dm^3 of sulphur dioxide at STP.**

Calculation

The formula of sulphur dioxide = SO_2

The relative atomic masses are; S = 32 and O = 16

$\therefore$ One mole of SO_2 = 32 + 2 (16) = 32 + 32 = 64 g

Molar volume of gas at STP = 22.4 dm^3

22.4 dm^3 at STP weighs 64 g

2.8 dm^3 .... $\frac{64}{1} \times \frac{2.8}{22.4} = 8$ g

2.8 dm^3 (STP) of SO_2 weighs 8 g.

Calculation of the volume of a gas from a chemical equation

Example : 3 g of pure carbon is completely burnt in oxygen.

Find the volume of carbon dioxide formed at STP.

Calculation $C + O_2 \longrightarrow CO_2$

1 mol 1 mol

1 mol of C = 12 g

$\therefore$ 3 g of C = $\frac{3g}{12g}$ = 0.25 mol of C

1 mol of C gives rise to 1 mol of CO_2

$\therefore$ 0.25 mol of C ... ... 0.25 mol of CO_2

But 1 mol of CO_2 occupies a volume of 22.4 dm^3 at STP

$\therefore$ 0.25 mol of CO_2 ... ... $\frac{22.4}{1} \times \frac{0.25}{1}$ = 5.6 dm^3 at STP

9.12 Determination of the Relative Molecular Mass of a Gas from its Molar Volume

The relative molecular mass of a gas can be experimentally found out by different methods. Some of these methods will be described in the later courses. In this section we shall describe how the relative molecular mass of a gas can be found out from its molar volume which can be determined by experiment. The procedure is as follows:

(1) The mass of a known volume of a gas at STP is experimentally determined.

(2) From these experimental results, the mass occupied by 22.4 dm^3 of the gas at STP is calculated.

(3) The calculated mass is equal to one mole of the gas and thus the numerical value of the calculated mass corresponds to the molecular mass of gas. We shall illustrate the procedure with an example in the following paragraphs.

By an experiment it was found that the volume of 1.4 dm^3 of oxygen gas at STP weighed 2 g.

1.4 dm^3 of oxygen gas at STP weighs 2 g

$\therefore$ 22.4 dm^3 ... ... $\frac{2}{1} \times \frac{22.4}{1.4}$ = 32 g

$\therefore$ One mole of oxygen gas = 32 g

$\therefore$ The relative molecular mass of oxygen = 32

How can we say that an oxygen molecule consists of two atoms of oxygen?

In Chapter 5 and 7 it has been described that each oxygen molecule consists of two atoms of oxygen. We can prove this in many ways and some of the methods will be described in the later courses. One such method will be described below.

The relative molecular mass divided by the relative atomic mass will give the number of atoms in the oxygen molecule.

For example, the relative molecular mass of oxygen as found out from its molar volume is 32.

The relative atomic mass of oxygen is 16.

$\therefore$ The number of oxygen atoms present in an oxygen molecule = $\frac{32}{16}$ = 2.

$\therefore$ One molecule of oxygen consists of 2 atoms of oxygen.

In the same way the relative molecular masses of hydrogen, nitrogen and chlorine can be found out from their molar volumes as described above. By dividing the relative molecular mass by the respective relative atomic mass, we can find the number of atoms present in one molecule of the gas.

The results indicate that one molecule of these gases consists of two atoms of the element.

Based upon these experimental results, we can now fully understand why the formula of oxygen is written as O_2, hydrogen as H_2, nitrogen as N_2 and chlorine as Cl_2.

SUMMARY

In this chapter, the measurement of the amount of substances in terms of number of atoms has been presented. The number of atoms present in the relative atomic mass in grams is the same for all elements. The number of molecules present in the relative molecular mass in grams is the same for all elements and compounds, i.e, the molecular mass in grams of all elements and compounds contain the same number of molecules. By experimental investigation, it was found that the numerical value of this number is 6.02×10^{23}. This number is known as the **Avogadro's Number** and is expressed by the symbol, L .The Avogadro's Number is constant for all elements, it is also called the **Avogadro's Constant** and may be defined by referring to the atomic mass in grams of any element. However, since ^{12}C has been chosen as the reference element, Avogadro's Number or constant is defined by referring to ^{12}C as follows: The Avogadro's number or constant is the number of carbon atom present in exactly 12 g of ^{12}C .One mole of a substance is the amount of substance which contains the same number of particles (atoms, molecules etc) as there are atoms in 12 g of ^{12}C.

One mole of any gas occupies a volume of 22.4 dm^3 at STP and this volume is called the molar volume of the gas. Avogadro's theory or law may be defined as "At the same temperature and pressure equal volume of all gases contain the same number of molecules".

Questions and Problems

1. Relative atomic masses of H= 1, Na = 23, Ca = 40, K = 39, Fe = 56, respectively. By assuming that in 1 g of hydrogen L numbers of hydrogen atoms are present, prove that the same number of atoms are present in the relative atomic mass in grams of the above elements.
2. H = 1, N = 14, O = 16, S = 32. Reason out that the same number of molecules are present in the relative molecular mass of H_2, N_2, O_2, H_2O and H_2S gases.
3. Explain by an example how chemists can count the atoms by weighing. What is the mass of each of the following?
   (a) 1 mole of sulphur atoms.
   (b) 0.5 mole of carbon atoms.
   (c) 0.25 mole of calcium atoms.
   (d) 1 mole of oxygen atoms.
   (e) 1 mole of oxygen molecules.
   (f) 1 mole of nitrogen atoms.
   (g) 0.5 mole of nitrogen molecules.
   (h) 0.25 mole of magnesium atoms.
   (i) 0.2 mole of water molecules.
   (j) 0.1 mole of carbon dioxide molecules.
   (k) 0.125 mole of sulphur dioxide molecules.
4. Convert the mass of each of the following to mole.
   (a) 32 g of oxygen atoms
   (b) 32 g of oxygen molecules.
   (c) 28 g of nitrogen atoms.
   (d) 14 g of nitrogen molecules.
   (e) 6 g of magnesium atoms.
   (f) 17 g of sulphur atoms.
   (g) 7 g of iron atoms.
   (h) 11 g of carbon dioxide molecules.
   (i) 3 g of water molecules.
   (j) 8 g of sulphur dioxide molecules.

5. Calculate the mass of each of the following. Use the relative atomic masses given in the table of relative atomic masses.
   (a) How many grams of potassium will contain the same number of atoms as 3 g of magnesium?
   (b) How many grams of carbon will contain the same number of atoms as 6 g of magnesium?
   (c) How many grams of oxygen, O_2 will contain the same number of molecules as 10 g of bromine, Br_2?
   (d) How many grams of copper will contain the same number of atoms as 8 g of sulphur?
   (e) How many grams of aluminium will contain the same number of atoms as there are molecules in 3 g of water?
6. Use the relative atomic masses given in the table.
   (a) How many moles of nitrogen, N_2 have the same number of particles as in 12 g of carbon?
   (b) How many moles of magnesium have the same number of particles as in 56 g of nitrogen, N_2?
   (c) How many moles of sulphur have the same number of particles as in 4 g of oxygen, O_2?
   (d) How many moles of calcium have the same number of atoms as in 6 g of magnesium?
   (e) How many moles of lithium have the same number of atoms as in 8 g of iron?
7. Give the equivalent in the moles for each of the following.
   (a) 3.2 dm^3 of nitrogen at STP.
   (b) 5.5 dm^3 of carbon dioxide at STP.
   (c) 11.2 dm^3 of oxygen at STP.
   (d) 1.6 dm^3 of chlorine at STP.
   (e) 1.4 dm^3 of hydrogen at STP.
   (f) 44.8 dm^3 of sulphur dioxide at STP.
8. What is the volume in dm^3 at STP of each of the following?
   (a) oxygen, O_2 0.5 mole
   (b) nitrogen N_2 0.2 mole
   (c) chlorine, Cl_2 0.4 mole
   (d) hydrogen, H_2 0.75 mole
   (e) carbon dioxide 0.4 mole
   (f) sulphur dioxide 0.5 mole
   (g) water vapour 0.9 mole
   (h) carbon dioxide 3 moles

9. What is the volume in dm^3 at STP of each of the following?

(a) 11 g of carbon dioxide
(b) 8 g of oxygen
(c) 7 g of nitrogen
(d) 35.5 g of chlorine
(e) 6.4 g of sulphur dioxide
(f) 4.5 g of water vapour

10. What is the mass of each of the following?

(a) 5.6 dm^3 of carbon dioxide at STP.
(b) 2.8 dm^3 of oxygen at STP.
(c) 11.2 dm^3 of nitrogen at STP.
(d) 4.48 dm^3 of chlorine at STP.
(e) 44.8 dm^3 of sulphur dioxide at STP.
(f) 2240 dm^3 of water vapour at STP.

11. 6 g of pure carbon is completely burnt in oxygen. The chemical equation for this reaction is:

$$C\,(s) + O_2\,(g) \longrightarrow CO_2\,(g)$$

(a) How many moles of carbon dioxide are formed?
(b) How many grams of carbon dioxide are formed? (C = 12)
(c) What is the volume in dm^3 at STP of carbon dioxide that forms in the reaction?
(d) How many moles of oxygen are used up?
(e) How many grams of oxygen are used up? (O = 16)
(f) How many dm^3 at STP of oxygen are used up?

12. A piece of pure sulphur is completely burnt in oxygen and 16 g of sulphur dioxide is formed. The chemical equation is:

$$S\,(s) + O_2\,(g) \longrightarrow SO_2\,(g)$$

(a) What is the volume at STP of SO_2 formed in the reaction?
(b) How many grams of sulphur are used up in the reaction? (S = 32)

13. By an experiment it was found that the mass of 5.6 dm^3 at STP of oxygen is 8g.

(a) What is the relative molecular mass of oxygen?
(b) Prove that one molecule of oxygen contains 2(atoms of oxygen. (O = 16)

14. The mass of 6 g of ozone gas occupied the volume of 2.8 dm^3 at STP.

(a) What is the relative molecular mass of ozone gas?
(b) How many atoms of oxygen are present in a molecule of ozone? (O = 16)

15. The volume of 3.2 dm^3 of nitrogen at STP weighs 4 g. What is the relative molecular mass of nitrogen and how many atoms of nitrogen are present in one molecule of nitrogen? (N =14)

16. The density of chlorine gas is 3.17 g dm^{-3} at STP. What is the relative molecular mass of chlorine and how many atoms of chlorine are present in one molecule? (Cl=35.5)

17. A gas consists of two elements carbon and hydrogen. The composition by weight of carbon and hydrogen is 6:0.5. The volume of 5.6 dm^3 of the gas at STP weighs 6.5 g.
    (a) Find the empirical formula of the gas.
    (b) Find the relative molecular mass of the gas.
    (c) Find the molecular formula of the gas.

18. Write TRUE or FALSE for each of the following statements.
    (a) The number of atoms present in the relative atomic mass in gram is the same for all elements.
    (b) One molecule of H_2O is 18 times lighter than 1 atom of H.
    (c) The Avogadro's Number or Constant is the number of carbon atom present in 12 g of ^{12}C.
    (d) One mole of oxygen atoms and one mole of oxygen molecules are similar in mass.
    (e) The density of a gas is usually expressed in grams per dm^3.

19. Fill in the blanks with a suitable word or words or symbol or unit or number.
    (a) The relative mass of one atom of carbon is ....
    (b) The Avogadro's Number can be represented by the symbol, L which is equivalent to ....
    (c) The volume of one mole of any gas at STP is ....
    (d) The mole ratio of Mg and S atom in the formation of MgS is ....
    (e) One mole of any gas occupies a volume of 22.4 dm^3 at STP and this volume of the gas is called .... of the gas.

20. Select the correct word or words given in the bracket.
    (a) One carbon atom is (10, 12, 14) times heavier than 1 atom of hydrogen.
    (b) The number of atoms present in the relative atomic mass in grams is (different, the same, equal) for all elements.
    (c) The Avogadro's Number or Constant is the number of carbon atoms present in exactly (12 g, 22 g, 32 g) of ^{12}C.
    (d) The volume of one mole of hydrogen at STP is (12.2 dm^3, 22.4 dm^3, 32.4 dm^3).
    (e) At the same temperature and pressure (similar, different, equal) volumes of all gases contain the same number of molecules.

21. Match each of the items given in list A with the appropriate item given in list B.

List A	List B
(a) Carbon burns in oxygen	(i) 2 atoms of oxygen
(b) One mole of any gas occupies a volume of 22.4 dm^3 at STP.	(ii) carbon dioxide is formed.
(c) One molecule of oxygen	(iii) 2 times heavier than 1 atom of H.
(d) The standard temperature and pressure	(iv) this volume is called the molar gas volume
(e) One molecule of H_2	(v) 0°C and 760 mmHg

22. Write TRUE or FALSE for each of the following statements.

(a) 18 g of liquid water occupies 22.4 dm^3 at STP.

(b) One molecule of hydrogen consists of 2 atoms of hydrogen. Therefore, the relative molcular mass of hydrogen is 2.

(c) If all gases have the same number of molecules, the number of moles are also the same.

(d) The symbol H represents one molecule of hydrogen.

(e) The mass of oxygen is 16 amu in a molecule of CO_2.

(f) The relative mass of one molecule of hydrogen is 1.

(g) One molecule of oxygen consists of 2 atoms of oxygen.

(h) One molecule of any gas at $0^{\circ}C$ and 760 mm Hg occupies a volume of 22.4 dm^3.

(i) The numerical value of Avogadro's Number is 6.02×10^{23}.

23. Fill in the blanks with a suitable word or words or symbol or unit or number.

(a) One mole of H_2 contains .... atoms.

(b) The mass of one mole of H atom is ....

(c) L is the symbol of Avogadro's Number and its numerical value is ....

(d) One mole of any gas occupies .... dm^3 at STP.

(e) The relative molecular mass of water is 18. Therefore, 18g of water contains .... number of water molecules .

(f) A group of particle containing 6.02×10^{23} is designated as ....

(g) 22.4 dm^3 of hydrogen gas at STP weighs ....g of hydrogen.

(h) Number of mole of 4g of hydrogen atom is ....

(i) The standard conditions for temperature and pressure chosen is ....

24. Answer the following questions.

(a) What is meant by the term STP ?

(b) What is relative molecular mass of Na_2CO_3 ? (Na =23. C = 12, O = 16)

(c) How many molecules are there in 1g of hydrogen ? (H = 1)

(d) Find the density of oxygen at STP ? (O = 16)

(e) How many hydrogen atoms are there in 1 mole of hydrogen molecules ? (H = 1)

(f) What is the volume of 1g of hydrogen at STP ?

(g) What is the mass of the carbon atom in one molecule of CO_2 ? (C = 12, O =16)

(h) What is the mass of 0.5 moles of water ? (H = 1, O = 16)

(i) How many atoms are there in 5 moles of oxygen atoms ? (O = 16)

(j) How many moles are there in 8 g of oxygen gas ? (O = 16)

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CHAPTER 10

OXYGEN AND ITS COMPOUNDS

10.1 Occurrence

Uncombined oxygen exists in the air, forming 23% by weight (or 21 % by volume) of the air. Oxygen in the combined state exists in water, sand or silica, silicates, and rocks.

10.2 Methods of Preparation of Oxygen

Laboratory preparation of oxygen from potassium chlorate

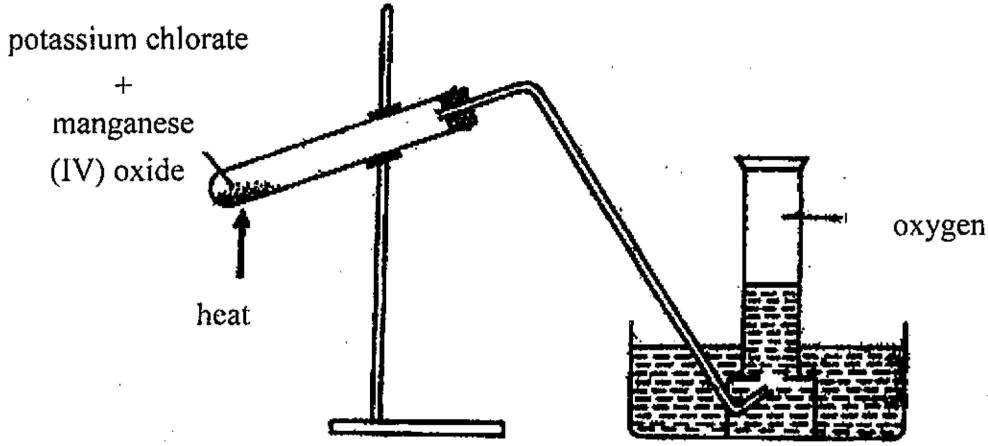

Fig. 10.1 Laboratory preparation of oxygen from potassium chlorate

∴ potassium chlorate	$\xrightarrow{\Delta}$	potassium chloride	+	oxygen
$2KClO_3$	$\xrightarrow{\Delta}$	$2KCl$	+	$3O_2$

A mixture of potassium chlorate and manganese (IV) oxide in the ratio of 4 : 1 by weight is placed in a hard glass tube. The apparatus is fitted up as shown in Fig. 10.1 and the hard glass tube is heated. Oxygen is evolved.

Since oxygen has about the same relative vapour density as air, and since it is only slightly soluble in water, it is collected by the downward displacement of water.

If required dry, it may be dried by passage through anhydrous calcium chloride, and collected in a syringe.

Test for oxygen

When a glowing splint of wood is put into the gas jar containing oxygen, it will be rekindled.

Note If potassium chlorate is heated alone, it gives off oxygen, but only at fairly high temperatures. If mixed with manganese (IV) oxide, the potassium chlorate gives off oxygen at a much lower temperature. On analysis of the residual mixture, it is found that the amount of manganese (IV) oxide is unchanged.

A chemical reaction that goes too slowly, may often be hastened by the addition of some substance that is found unchanged after the reaction is over. When such a substance is used in a reaction for this purpose, it is said to catalyse the reaction. The substance itself is called a catalyst, and the process is called catalysis.

Industrial preparation

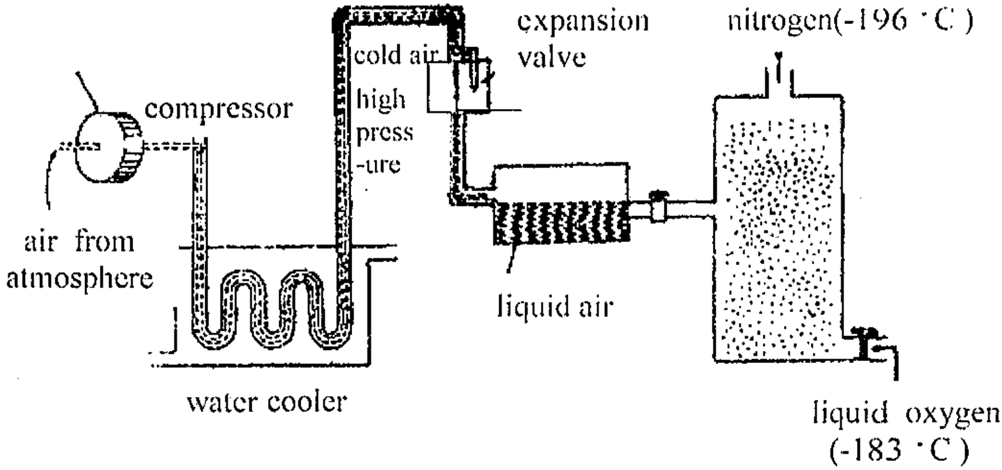

Fig. 10.2 Industrial preparation of oxygen

The best process for obtaining oxygen industrially is from liquid air.

Air is first cooled by allowing compressed air to expand through a valve (Fig. 10.2). This cool air is recycled and is again recompressed and expanded through the valve. This results in further cooling. When these successive coolings are finally sufficient (i.e., temperature is reduced lower than - 196°C), the air is liquefied. This liquefied air is then fed into the fractionating column. The lower boiling nitrogen (b.p. - 196°C) will come out from the top of the fractionating column and the higher boiling oxygen (b.p,- 183°C) will come out from the bottom of the column.

10.3 The Activity Series

K
Na
Ca
Mg
Al
Zn
Fe
Pb
H
Cu
Ag

Electropositivity decreases from top to bottom.

By studying the action of air, water and acids on the metals and the ease by which the metal oxides may be reduced to metal, we may arrange the metals in a series known as the Activity Series. A summarized form of the activity series outlining some trends in the reactivity of the metals is given in Appendix 4. It may be seen that with few exceptions there are very definite trends as the series is descended. Reference will be made to this series in the appropriate sections and the reasons for the trends will be discussed.

10.4 Properties of Oxygen

1. Oxygen is a colourless gas, without taste or smell. It is only slightly soluble in water and has approximately the same relative vapour density as air.
2. Oxygen will not burn, but it supports combustion.
3. Many metals and non-metals burn in oxygen, forming oxides.

Action with metals

The readiness with which oxygen reacts with metals is summarized in the following list arranged according to the activity series.

K Na Ca Mg Al Fe Pb Cu	When these metals are heated in air, the oxygen reacts with potassium most readily, and with copper least readily.
Hg Ag Au	These metals react with oxygen at high temperature.

Example :

A burning piece of magnesium ribbon continues burning in oxygen with a dazzling white flame, leaving a white powder as residue. This residue is magnesium oxide.

magnesium + oxygen $\longrightarrow$ magnesium oxide

$$2Mg + O_2 \longrightarrow 2MgO$$

A red hot iron wire burns in oxygen with a yellowish flame giving off showers of golden sparks.

iron + oxygen $\longrightarrow$ tri iron tetroxide

$$3\,Fe + 2O_2 \longrightarrow Fe_3O_4$$

A heated piece of sodium metal burns in oxygen with a yellowish flame to form a solid which dissolves in water. The solution obtained turns red litmus blue.

sodium + oxygen $\longrightarrow$ sodium oxide

$$4Na + O_2 \longrightarrow 2Na_2O$$

sodium oxide + water $\longrightarrow$ sodium hydroxide

$$Na_2O + H_2O \longrightarrow 2NaOH$$

Action with non-metals

Non-metals such as phosphorus, sulphur and carbon react with oxygen to form oxides.

Example :

A small piece of heated phosphorus (only red phosphorus may be used) burns in oxygen giving off white fumes which consists of oxides of phosphorus.

phosphorus	+	oxygen	$\longrightarrow$	phosphorus (V) oxide
$4P$	+	$5O_2$	$\longrightarrow$	P_4O_{10}
phosphorus	+	oxygen	$\longrightarrow$	phosphorus (III) oxide
$4P$	+	$3O_2$	$\longrightarrow$	P_4O_6

These oxides dissolve in water to give solutions which turn blue litmus red.

water	+	phosphorus (V)oxide	$\longrightarrow$	phosphoric acid
$6H_2O$	+	P_4O_{10}	$\longrightarrow$	$4H_3PO_4$
water	+	phosphorus (III) oxide	$\longrightarrow$	phosphorus acid
$6H_2O$	+	P_4O_6	$\longrightarrow$	$4H_3PO_3$

Burning sulphur continues to burn in oxygen to form misty, pungent fumes which are sulphur dioxide.

sulphur	+	oxygen	$\longrightarrow$	sulphur dioxide
S	+	O_2	$\longrightarrow$	SO_2

This gas dissolves in water to give sulphurous acid. The acid solution turns blue litmus red.

water	+	sulphur dioxide	$\longrightarrow$	sulphurous acid
H_2O	+	SO_2	$\longrightarrow$	H_2SO_3

When a burning piece of charcoal is put into a gas jar of oxygen, it burns and emits a shower of sparks to form a colourless gas, carbon dioxide.

carbon	+	oxygen	$\longrightarrow$	carbon dioxide
C	+	O_2	$\longrightarrow$	CO_2

Carbon dioxide gas dissolves in water to form carbonic acid. This is a very weak acid and turns blue litmus pink but not red.

water	+	carbon dioxide	$\longrightarrow$	carbonic acid
H_2O	+	CO_2	$\longrightarrow$	H_2CO_3

When silicon is burnt in oxygen, silicon dioxide (silica) is obtained.

silicon	+	oxygen	$\longrightarrow$	silicon dioxide
Si	+	O_2	$\longrightarrow$	SiO_2

Sand is one form of silica.

Since silicon dioxide is insoluble in water, no solution is formed and there will be no reaction with litmus paper. However, the reactions with alkalis show that silicon dioxide has acidic character.

silicon dioxide	+	sodium hydroxide	$\longrightarrow$	sodium silicate	+	water
SiO_2	+	$2NaOH$	$\longrightarrow$	Na_2SiO_3	+	H_2O

Note The combination of a substance with oxygen is called oxidation.

10.5 Uses of Oxygen

Oxygen is used as an aid to breathing where the natural supply of oxygen is insufficient. It is also employed as a mixture with hydrogen or acetylene for the production of high temperatures. The oxyacetylene flame is used for cutting and welding steel. Oxygen also finds application in the other industries such as the manufacture of linoleum, varnishes, steel etc.

10.6 Ozone O_3

Oxygen and ozone are allotropes (section 15.2) of the same element. The difference between them is that oxygen is a diatomic molecule (O_2) and ozone is a triatomic molecule (O_3).

Preparation of ozone

Ozone is prepared by the passage of a silent electric discharge through oxygen.

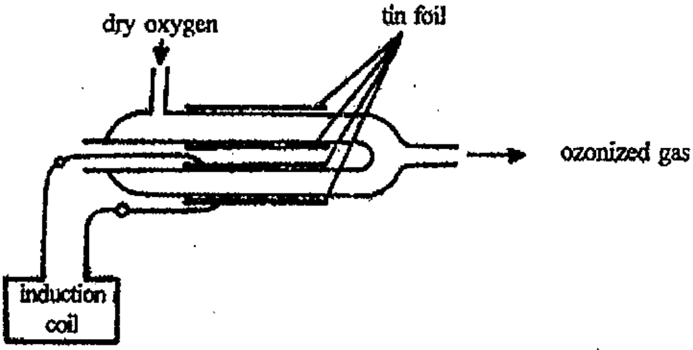

Fig. 10.3 Preparation of ozone gas

oxygen	$\longrightarrow$	ozone
$3O_2$	$\longrightarrow$	$2O_3$

The apparatus is set up as shown in Fig. 10.3. Dry oxygen is passed through the space between the two glass tubes. Each tube is coated with tin foil which serves as electrodes and is connected to the terminals of an induction coil. The silent electrical discharge passing through the oxygen from one electrode to the other causes partial conversion of the oxygen to ozone. This ozonized oxygen should not be allowed to come into contact with rubber, which is attacked by ozone.

Test for ozone

Ozone possesses a fishy smell. When mercury makes contact with ozone, the mercury 'tail' is observed, i.e., mercury sticks to the glass as it flows across, leaving a trail of mercury.

Properties of ozone

1. Ozone gas has a fishy smell.
2. It oxidizes;

(a) lead (II) sulphide to lead (II) sulphate.

lead(II) sulphide (black) + ozone $\longrightarrow$ lead (II) sulphate (white) + oxygen

$$PbS + 4O_3 \longrightarrow PbSO_4 + 4O_2$$

(b) hydrogen sulphide to sulphuric acid.

hydrogen sulphide + ozone $\longrightarrow$ sulphuric acid + oxygen

$$H_2S + 4O_3 \longrightarrow H_2SO_4 + 4O_2$$

3. It also liberates iodine from potassium iodide in acidic solution.

potassium iodide (colourless) + sulphuric acid + ozone $\longrightarrow$ iodine (brown) + oxygen + potassium sulphate + water

$$2KI + H_2SO_4 + O_3 \longrightarrow I_2 + O_2 + K_2SO_4 + H_2O$$

Uses of ozone

Ozone is used in ventilation systems to purify air. It is also used to sterilize water.

10.6.1 Ozone depletion

Chlorofluorocarbons (CFCs) are widely used as coolants in refrigerators and car air conditioners, as solvents for cleaning computer and circuit boards, and in foam packaging and insulation. Evidence indicates that the chemicals are finding their way to the stratosphere and destroying the ozone there. The ozone layer filters out most of the ultraviolet radiation in sunlight. It is believed that destruction of the layer would cause a dramatic increase in skin cancer cases, a drop in crop yields, and reduction of the microscopic organisms in the food chain of the oceans.

Since negotiation of the Montreal Protocol, scientists have found that the ozone in the stratosphere is being depleted more rapidly than had been thought. Measurements reported recently showed greatly increased levels of biologically damaging ultraviolet radiation at the earth's surface in the Antarctic, where a "hole" in the ozone layer has periodically appeared.

There was considerable controversy recently over the effects of aerosol propellants, particularly the chlorofluorocarbons, on the ozone shield in the earth's stratosphere. The two most widely used chlorofluorocarbons are chlorofluorocarbon 11 ($CFCl_3$) and chlorofluorocarbon 12 (CF_2Cl_2), often referred to by their trade names, Freon 11 and Freon 12. Chlorofluorocarbon 11 is the propellant most used in aerosol spray cans for items other than food (paint, polish, hair spray, and deodorants, for example), and chlorofluorocarbon 12 is used primarily as a circulating fluid in refrigerators and air conditioners.

Ozone (O_3) is a form of elementary oxygen produced in the upper atmosphere (the stratosphere layer 6-30 miles above the surface). Ultraviolet radiation from the sun converts ordinary oxygen (O_2) into ozone by the reactions

$$O_2 + h\upsilon \rightarrow 2O \quad \text{and}$$

$$O + O_2 \rightarrow O_3,$$

where h_2 is the symbol for radiation and O stands for a single oxygen atom.

Once created, the ozone in the stratosphere is a good absorber of ultraviolet radiation. Without ozone, much of this radiation would reach the earth, raising surface temperatures and increasing the frequency of nonmelanistic skin cancer among people regularly exposed to the sun. But ozone is only moderately stable. Not only does it easily decompose back to ordinary oxygen, but more importantly, it reacts with components of the atmosphere, such as nitric oxide (NO) and chlorine (Cl) atoms. These reactions are the source of the current problem.

Chlorofluorocarbons 11 and 12 are low molecular weight compounds and, once released into the atmosphere, can diffuse to the stratosphere. There they are acted upon by ultraviolet radiation, according to the reaction (using chlorofluorocarbon 11 as an example)

$$CFCl_3 + h\upsilon \rightarrow CFCl_2 + Cl$$

The chlorine atom thus produced can react with ozone

$$(Cl + O_3 \rightarrow ClO + O_2),$$

and the chlorine oxide may react with an oxygen atom

$$(ClO + O \rightarrow Cl + O_2)$$

to regenerate the original chlorine atom. This sequence is an example of a chain reaction in which the original event, destruction of the ozone molecule, can be repeated hundreds or thousands of times for every chlorine atom that is formed by the action of ultraviolet radiation.

Since one chlorofluorocarbon molecule in the stratosphere has the potential to destroy many ozone molecules and since nearly 1 million tons of chlorofluorocarbons 11 and 12 are produced annually, the amount of ozone that could possibly be destroyed by chlorofluorocarbons is enormous.

10.7 Hydrogen Compounds of Oxygen

When hydrogen combines with one other element only, the compound formed is called a hydride. Hydrogen combines with oxygen to form two different hydrides, namely, water, H_2O and hydrogen peroxide, H_2O_2.

Preparation of a solution of hydrogen peroxide H_2O_2

Hydrogen peroxide may be prepared by the action of certain metallic peroxides with acids. The materials usually used are barium peroxide and dilute sulphuric acid. The barium sulphate produced in the reaction is insoluble and can be filtered off. The filtrate is an aqueous solution of hydrogen peroxide.

barium peroxide	+	sulphuric acid (dil)	$\longrightarrow$	barium sulphate	+	hydrogen peroxide
BaO_2	+	H_2SO_4	$\longrightarrow$	$BaSO_4$	+	H_2O_2

Properties of hydrogen peroxide

1. The pure compound is a syrupy liquid. It is usually used as a dilute solution in water. Hydrogen peroxide solution is sold retail as "10 volume" or "20 volume" solution, i.e., at STP (standard temperature and pressure) 10 cm^3 or 20 cm^3 of oxygen gas is liberated by heating 1 cm^3 of that hydrogen peroxide solution.

2. Evolution of oxygen occurs when hydrogen peroxide solution is heated.

hydrogen peroxide	$\longrightarrow$	water	+	oxygen
$2H_2O_2$	$\longrightarrow$	$2H_2O$	+	O_2

3. Hydrogen peroxide oxidizes lead (II) sulphide to lead (II) sulphate.

lead (II) sulphide (black) + hydrogen peroxide ⟶ lead (II) sulphate (white) + water

$$PbS + 4H_2O_2 \longrightarrow PbSO_4 + 4H_2O$$

4. When hydrogen peroxide reacts with acidified potassium iodide solution, iodine is produced.

potassium iodide + sulphuric acid + hydrogen peroxide ⟶ potassium sulphate + iodine (brown) + water

$$2KI + H_2SO_4 + H_2O_2 \longrightarrow K_2SO_4 + I_2 + 2H_2O$$

5. Hydrogen peroxide reacts with some compounds to give oxygen.

lead (IV) oxide (brown) + hydrogen peroxide ⟶ lead(II) oxide (yellow) + water + oxygen

$$PbO_2 + H_2O_2 \longrightarrow PbO + H_2O + O_2$$

silver oxide + hydrogen peroxide ⟶ silver + water + oxygen

$$Ag_2O + H_2O_2 \longrightarrow 2Ag + H_2O + O_2$$

potassium permanganate (violet) + sulphuric acid + hydrogen peroxide ⟶ manganese(II) sulphate (white) + potassium sulphate + water + oxygen

$$2KMnO_4 + 3H_2SO_4 + 5H_2O_2 \longrightarrow 2MnSO_4 + K_2SO_4 + 8H_2O + 5O_2$$

ozone + hydrogen peroxide ⟶ water + oxygen

$$O_3 + H_2O_2 \longrightarrow H_2O + 2O_2$$

6. The decomposition of hydrogen peroxide is catalysed by many substances. (e.g., manganese (IV) oxide, finely powdered gold, platinum)

hydrogen peroxide ⟶ water + oxygen

$$2H_2O_2 \longrightarrow 2H_2O + O_2$$

Uses of hydrogen peroxide

Hydrogen peroxide is used in restoring pictures. Hydrogen sulphide in the air slowly reacts with the white lead pigment (lead (II) carbonate) of the paint to produce lead (II) sulphide, which is brown and makes the picture dingy. Washing with hydrogen peroxide restores the white colour. Hydrogen peroxide is used commercially for the bleaching of paper pulp, cotton and other natural fibers.

SUMMARY

Uncombined oxygen exists in air, forming 21% by volume of the air. Oxygen in the combined state exists in water, sand or silica, silicates, and rocks.

Oxygen can be prepared in laboratory from potassium chlorate by heating. The best process for obtaining oxygen industrially is from liquid air. Oxygen is a colorless gas without taste or smell and only slightly soluble in water. It has some relative vapour density as air. It supports burning. So many metals burn in oxygen, forming oxides.

Oxygen is used as an aid to breathing and also in the manufacture of linoleum, varnishes, steel, etc. The oxyacetylene flame is used for cutting and welding steel.

Oxygen and ozone are allotropes of the same element .Ozone is used as in ventilation systems to purify air .It is also used to sterilize water.

CFCs (chlorofluorocarbons) are widely used as coolants in refrigerators and car air conditioners, as solvents for cleaning computer and circuit boards and in foam packing and insulation. The ozone layer filters out most of the ultraviolet radiation in sunlight. Since one chlorofluorocarbon molecule in the stratosphere and since negotiation protocol, scientist found that the ozone in the atmosphere is being depleted more rapidly than had been thought. Instead of CFCs, hydrochlorofluorocarbons (HCFCs) have been developed to reduce the ozone depletion.Hydrogen combines with oxygen to form two different hydrides, hydrogen peroxide, H_2O_2 and water, H_2O. Hydrogen peroxide is used commercially for the bleaching of paper pulp, cotton and other natural fibers.

Questions and Problems

1. (a) Outline a method by which oxygen is manufactured from air.

   (b) Name two compounds which, when heated alone, give off oxygen as the only gaseous product, and write equations for their decomposition.

2 Write equations for the reactions of each of the following substance with excess of oxygen.

   (a) hydrogen (b) magnesium (c) carbon (d) zinc (e) silicon

Write equations for the reactions, if any, of these products with

(i) dilute hydrochloric acid (ii) sodium hydroxide

3. On analysis of a compound obtained from burning iron in oxygen, it is found that the constituents of this compound are 167.7 g of iron and 48 g of oxygen. What is the empirical formula of this compound? (Fe = 55.9, O = 16)

4. Calculate the mass of magnesium oxide which is obtained on burning 4 g of magnesium in the air. (Mg = 24, O = 16)

5. Calculate the mass of magnesium which reacts with 1 g of oxygen.

6. When a mixture of 4 g of magnesium and magnesium oxide was heated in air, 5 g of magnesium oxide was obtained. Calculate the percentage by mass of magnesium and magnesium oxide in mixture. (Mg = 24, O = 16)

7. Write equations for the reactions, if any, of the following:

   (a) a piece of gold in air

   (b) mercury and oxygen were heated vigorously

   (c) silicon dioxide mixed with water

   (d) dissolving sodium oxide in water.

8. How can oxygen be converted into its allotropic form? Give equations for these reactions in which ozone reacts with compounds to liberate oxygen.

9. What would you observe in the following reactions? Write equations for the reactions.

   (a) When hydrogen peroxide solution is added into acidified potassium permanganate solution.
   (b) When hydrogen peroxide solution is heated.
   (c) Ozone gas is passed into potassium iodide solution.
   (d) When hydrogen peroxide solution is poured over manganese (IV) oxide.

10. Write TRUE or FALSE for each of the following statements.

   (a) Oxygen will burn but it cannot support combustion.

   (b) Burning sulphur continues to burn in oxygen to form sulphur dioxide.

   (c) Ozone possesses sweet smell.

   (d) Hydrogen peroxide oxidizes lead (II) sulphide to lead (II) sulphate.

   (e) Heat is needed in the laboratory preparation of oxygen from hydrogen peroxide.

   (f) Oxygen is lighter than air.

   (g) Oxygen and ozone are allotropes of the same element.

   (h) Oxygen will burn but it cannot support combustion.

11. Fill in the blanks with a suitable word or words.

(a) Carbon dioxide dissolves in water to form ....

(b) Many metals and non-metals burn in oxygen, forming ....

(c) When silicon is burnt in oxygen, ... is obtained.

(d) Ozone gas has a ....smell.

(e) Hydrogen peroxide may be prepared by the action of metallic peroxides with....................

(f) The combination of a substance with oxygen is called ....

(g) Ozone is used in ventilation system to purify air and to .... water.

(h) Many metals and non-metals burn in oxygen, forming ....

12. Answer the following questions are base on ozone depletion.

(a) What is CFCs?

(b) Name the three elements present in CFCs in connection with ozone depletion.

(c) Which of these elements is the most reactive?

(d) If the ozone (O_3) reacted with the most reactive element in CFCs, what two products would be formed?

(e) What are the uses of CFCs?

(f) Where is ozone layer found?

(g) What does ozone layer filter out?

(h) Give one for the harmful radiation in the sunlight.

(i) How ordinary oxygen is converted into ozone?

(j) How CFCs destroy ozone layer?

(k) Estimate amount of chlorofluorocarbon 11 and 12 produced annually.

(l) How can you protect the ozone hole problem ?

13. Select the correct word or words given in the brackets.

(a) If required dry, it may be dried by passage through anhydrous (calcium oxide, calcium chloride, calcium nitrate) and collected in a syring.

(b) Many (oxides, hydroxides, metals) and non-metals burn in oxygen, forming oxides.

(c) The combination of a substance with (nitrogen, hydrogen, oxygen) is called oxidation.

(d) Oxygen is used as an aid to (heating, breathing, eating) where a natural supply of oxygen is insufficient.

(e) Ozone is used in ventilation system to (sterilize water, purify air, purify soil).

(f) (Oxygen, Ozone, Hydrogen) is used as an oxyacetylene flame in the cutting and welding of steel.

(g) Oxygen is collected by the (upward, downward, under) displacement of water .

(h) (Metals, Non-metals, Salts) react with oxygen to give oxides.

(i) A red hot iron wire burns in oxygen to form (iron (II) oxide, iron (III) oxide, tri iron tetroxide).

14. Match each of the items given in list A with the appropriate item given in list B.

List A	List B
(a) A mixture of potassium chlorate and manganese (IV) oxide is heated in a hard glass tube	(i) The process is called catalysis
(b) The substance itself is called a catalyst.	(ii) and to sterilize water.
(c) A burning magnesium ribbon placed in oxygen	(iii) carbonic acid is formed.
(d) Carbon dioxide gas dissolves in water	(iv) magnesium oxide is formed.
(e) Ozone is used in ventilation system to purify air.	(v) oxygen is evolved.

15. Answer the following questions.

(a) Write down the activity series.

(b) What is the main gas in the air ?

(c) Give three uses of oxygen gas.

(d) What are the three things needed for combustion or burning ?

(e) Write equations in symbols for the following reactions.

(i) Red hot iron wire burns in oxygen.

(ii) Red hot phosphorus burns in excess oxygen.

(f) Which method can be used to manufacture oxygen gas ?

Write down the name of the important steps of the process.

(g) What will happen when mercury and oxygen were heated vigorously?

(h) How would you prepare oxygen gas in the laboratory?

(i) Describe the laboratory preparation of dry oxygen from potassium chlorate.

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CHAPTER 11

OXIDES AND HYDROXIDES

11.1 Oxides

An oxide is a compound which contains only oxygen and one other element.

There are many kinds of oxides. The main types of oxides are ;

(1) basic oxides

(2) acidic oxides

(3) amphoteric oxides

(4) neutral oxides

(5) peroxides

(6) compound oxides

Basic oxides

A basic oxide is a metallic oxide. It is formed by the reaction of metals with oxygen.

metals	+	**oxygen**	$\longrightarrow$	**metallic oxides** (most of which are basic in character)

Properties

1. Solubility in water

Some basic oxides are soluble in water but some are not. (Refer to activity series) e.g., Na_2O, K_2O, are soluble oxides, while MgO, CuO, Ag_2O are not.

Soluble basic oxides react with water, forming hydroxide solutions (alkalis). These solutions turn red litmus blue.

sodium oxide	+	water	$\longrightarrow$	sodium hydroxide
Na_2O	+	H_2O	$\longrightarrow$	$2NaOH$

Other examples

basic oxide	alkali
K_2O (soluble in water)	KOH
CaO (slightly soluble in water)	$Ca(OH)_2$

2. Action with acids

A basic oxide reacts with an acid to produce a salt and water only.

(a) Soluble oxide

sodium oxide + hydrochloric acid (dil) $\longrightarrow$ sodium chloride + water

$$Na_2O + 2HCl \longrightarrow 2NaCl + H_2O$$

(b) Insoluble oxide

copper (II) oxide + sulphuric acid (dil) $\longrightarrow$ copper (II) sulphate + water

$$CuO + H_2SO_4 \longrightarrow CuSO_4 + H_2O$$

Properties of basic oxides can be summarized as follows :

$\left.\begin{array}{l}\text{K}\\ \text{Na}\\ \text{Ca}\end{array}\right\}$ Oxides of these metals are soluble in water forming alkalis

$\left.\begin{array}{l}\text{Mg}\\ \text{Al}\\ \text{Zn}\\ \text{Fe}\\ \text{Pb}\\ \text{Cu}\end{array}\right\}$ Oxides of these metals (except Al) can be made from the metal by the action of nitric acid and then decomposing by heating (i.e., by heating the nitrates of the metals)

K to Zn: Oxides of these metals are not reduced to metal by hydrogen

$\left.\begin{array}{l}\text{Hg}\\ \text{Ag}\\ \text{Au}\end{array}\right\}$ Oxides of these metals decompose when heated

Acidic oxides

An acidic oxide is a non-metallic oxide. It is formed by the reaction of non-metals with oxygen.

non-metals	+	**oxygen**	$\longrightarrow$	**non-metallic oxides** (most of which are acidic in character)

Properties

1. Solubility in water

Some acidic oxides are soluble in water but some are not. e.g., CO_2, SO_2, SO_3, P_4O_{10} are soluble oxides while SiO_2, is an insoluble oxide.

Soluble acidic oxides dissolve in water to form acid solutions. These solutions turn blue litmus red.

sulphur trioxide	+	water	$\longrightarrow$	sulphuric acid
SO_3	+	H_2O	$\longrightarrow$	H_2SO_4
phosphorus (V) oxide	+	water	$\longrightarrow$	phosphoric acid
P_4O_{10}	+	$6H_2O$	$\longrightarrow$	$4H_3PO_4$

Other examples

acidic oxide	**acid**
CO_2	H_2CO_3 (carbonic acid)
SO_2	H_2SO_3 (sulphurous acid)

2. Action with basic oxides

Soluble acidic oxides react with basic oxides to give normal salts (without hydrogen).

carbon dioxide	+	sodium oxide	$\longrightarrow$	sodium carbonate
CO_2	+	Na_2O	$\longrightarrow$	Na_2CO_3

3. Action with alkalis

(a) Soluble acidic oxide

Soluble acidic oxides react with alkali solutions to give either normal salts and water, or an acid salt.

With limited amount of acidic oxide, salt and water are formed.

carbon dioxide + calcium hydroxide $\longrightarrow$ calcium carbonate + water

$$CO_2 + Ca(OH)_2 \longrightarrow CaCO_3 + H_2O$$

With in excess amount of acidic oxide, an acid salt is formed.

carbon dioxide + calcium hydroxide $\longrightarrow$ calcium hydrogencarbonate

$$2CO_2 + Ca(OH)_2 \longrightarrow Ca(HCO_3)_2$$

(b) Insoluble acidic oxides

Insoluble acidic oxides do not react with water. They react with basic oxides and alkalis.

With basic oxide

silicon dioxide + sodium oxide $\longrightarrow$ sodium silicate

$$SiO_2 + Na_2O \longrightarrow Na_2SiO_3$$

With alkali

silicon dioxide + sodium hydroxide $\longrightarrow$ sodium silicate + water

$$SiO_2 + 2NaOH \longrightarrow Na_2SiO_3 + H_2O$$

Amphoteric oxides

An amphoteric oxide is a metallic oxide which possesses both basic and acidic properties, i.e., it can react with both acid and alkali.

Example : ZnO and Al_2O_3

(a) Basic property (reaction with acids)

zinc oxide + sulpuric acid (dil) $\longrightarrow$ zinc sulphate + water

$$ZnO + H_2SO_4 \longrightarrow ZnSO_4 + H_2O$$

aluminium oxide + hydrochloric acid $\longrightarrow$ aluminium chloride + water

$$Al_2O_3 + 6HCl \longrightarrow 2AlCl_3 + 3H_2O$$

(b) Acidic property (reaction with alkalis)

zinc oxide + sodium hydroxide + water $\longrightarrow$ sodium zincate

$$ZnO + 2NaOH + H_2O \longrightarrow Na_2ZnO_2.2H_2O\text{(or)}Na_2Zn(OH)_4$$

aluminium oxide + sodium hydroxide + water $\longrightarrow$ sodium aluminate

$$Al_2O_3 + 2NaOH + 3H_2O \longrightarrow 2NaAl(OH)_4$$

Neutral oxides

A neutral oxide is an oxide which shows neither basic nor acidic character. e.g., carbon monoxide, CO, dinitrogen oxide, N_2O. (Although water is neutral to litmus, it is not a neutral oxide.)

Peroxides

Those oxides that react with acid to give salt and hydrogen peroxide are called peroxides.

barium peroxide	+	sulphuric acid	$\longrightarrow$	barium sulphate	+	hydrogen peroxide
BaO_2	+	H_2SO_4	$\longrightarrow$	$BaSO_4$	+	H_2O_2

PbO_2 lead (IV) oxide, MnO_2 manganese (IV) oxide and NO_2 nitrogen dioxide are not peroxides. They do not give hydrogen peroxide on reaction with acids.

Compound oxides

A compound oxide is an oxide, formed by the combination of two different oxides of the same element.

Example : Di lead (II) lead (IV) oxide (red lead oxide) Pb_3O_4 is a compound consisting of lead (II) oxide and lead (IV) oxide. Pb_3O_4can be written as $2PbO\ .\ PbO_2$.

The following reaction supports this statement.

red lead oxide	+	nitric acid	$\longrightarrow$	lead (II) nitrate	+	water	+	lead (IV) oxide
$Pb_3\,O_4$	+	$4HNO_3$	$\longrightarrow$	$2Pb(NO_3)_2$	+	$2H_2O$	+	PbO_2

Compare this reaction with the following :

lead(II) oxide	+	lead (IV) oxide	+	nitric acid	$\longrightarrow$	lead(II) nitrate	+	water	+	lead (IV) oxide
$2PbO$	+	PbO_2	+	$4HNO_3$	$\longrightarrow$	$2Pb(NO_3)_2$	+	$2H_2O$	+	PbO_2
base				acid		salt		water		

Pb_3O_4 unchanged

Other examples : Fe_3O_4 ($FeO.\ Fe_2O_3$)

Mn_3O_4 ($2MnO.\ MnO_2$)

11.2. Preparation of Oxides

Several methods are used for the preparation of oxides.

Method 1 Direct combination of an element with oxygen

Metals or non-metals react with oxygen to give oxides.

metal + oxygen $\longrightarrow$ metallic oxide (basic oxide)

non-metal + oxygen $\longrightarrow$ non-metallic oxide (acidic oxide)

Example :

magnesium + oxygen $\longrightarrow$ magnesium oxide

$$2Mg + O_2 \longrightarrow 2MgO$$

carbon + oxygen $\longrightarrow$ carbon dioxide

$$C + O_2 \longrightarrow CO_2$$

sodium + oxygen $\longrightarrow$ sodium oxide

$$4Na + O_2 \longrightarrow 2Na_2O$$

sodium + oxygen (excess) $\longrightarrow$ sodium peroxide

$$2Na + O_2 \longrightarrow Na_2O_2$$

Method 2 Decomposition of nitrate

Most heavy metallic nitrates decompose on heating. They decompose to metallic oxide, nitrogen dioxide and oxygen.

Example :

zinc nitrate $\xrightarrow{\Delta}$ zinc oxide + nitrogen dioxide + oxygen

$$2Zn(NO_3)_2 \xrightarrow{\Delta} 2ZnO + 4NO_2 + O_2$$

copper(II) nitrate $\xrightarrow{\Delta}$ copper (II)oxide + nitrogen dioxide + oxygen

$$2Cu(NO_3)_2 \xrightarrow{\Delta} 2CuO + 4NO_2 + O_2$$

Method 3 Decomposition of carbonates

Most of the carbonates decompose to oxides and carbon dioxide on heating.

Example :

calcium carbonate $\xrightarrow{\Delta}$ calcium oxide + carbon dioxide

$$CaCO_3 \xrightarrow{\Delta} CaO + CO_2$$

Method 4   Decomposition of hydroxides

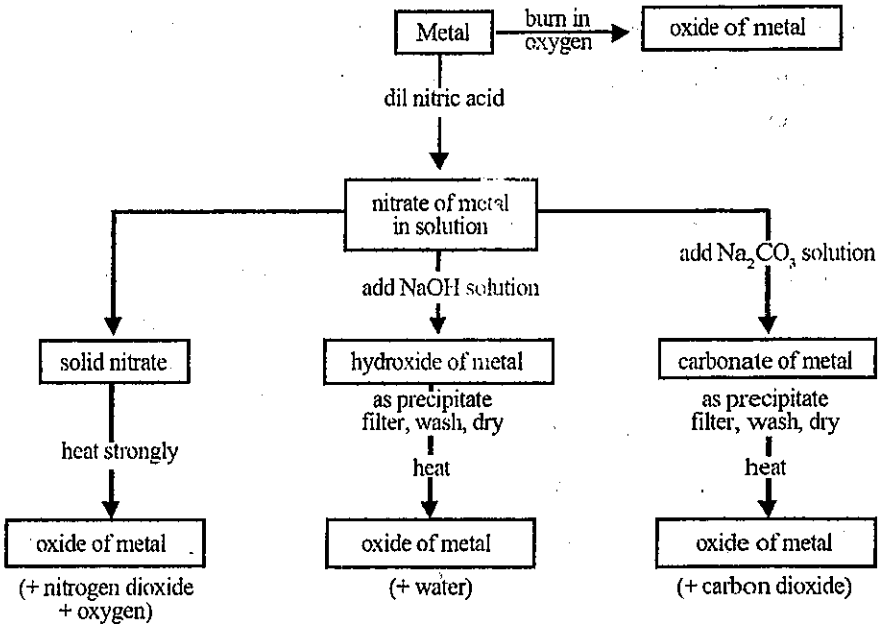
Some metallic hydroxides decompose to oxides and water on heating.

Example :

iron (III) hydroxide $\xrightarrow{\Delta}$ iron (III) oxide + water

$2Fe(OH)_3 \xrightarrow{\Delta} Fe_2O_3 + 3H_2O$

11.3. Summary Preparative Methods for Obtaining Oxides from Metals or their Soluble Salts



11.4 Hydroxides

Hydroxides can be classified as

(i) basic hydroxides (eg. NaOH, KOH, etc---)

(ii) amphoteric hydroxides (eg. $Zn(OH)_2$, $Al(OH)_3$)

(iii) acidic hydroxides (eg. $(HO)_2SO_2$, $HONO_2$)

Hydroxides which are soluble in water are called alkalis.

11.5 Preparation of Hydroxides

Soluble hydroxides of metals

Hydroxides of metals which are soluble in water may be prepared by the following methods.

Method 1 Action of metal with water

Some metals react with water to give soluble hydroxides of metals and hydrogen.

$$\text{sodium} + \text{water} \longrightarrow \text{sodium hydroxide} + \text{hydrogen}$$

$$2Na + 2H_2O \longrightarrow 2NaOH + H_2$$

Method 2 By dissolving soluble metal oxides in water

Some soluble metal oxides, not only dissolve in water, but react with water to give hydroxides.

$$\text{sodium oxide} + \text{water} \longrightarrow \text{sodium hydroxide}$$

$$Na_2O + H_2O \longrightarrow 2NaOH$$

$$\text{sodium peroxide} + \text{water} \longrightarrow \text{sodium hydroxide} + \text{nascent oxygen}$$

$$Na_2O_2 + H_2O \longrightarrow 2NaOH + (O)$$

Note Oxygen in the atomic state is called nascent oxygen. Two atoms of nascent oxygen combine to form molecular oxygen.

Insoluble hydroxides of metals

Insoluble hydroxides of metals are not prepared by the above methods. They may be prepared by using the following method.

Method : By the action of soluble salt with alkali

Addition of soluble alkalis (NaOH, KOH) to soluble salts of heavy metals give insoluble hydroxides.

$$\textbf{soluble salt of heavy metal} + \textbf{alkali} \longrightarrow \textbf{insoluble hydroxide of metal} + \textbf{salt}$$

Example :

$$\text{copper (II) nitrate} + \text{potassium hydroxide} \longrightarrow \text{copper (II) hydroxide} + \text{potassium nitrate}$$

$$Cu(NO_3)_2 + 2KOH \longrightarrow Cu(OH)_2 + 2KNO_3$$

11.6 Properties of Hydroxides

The properties of hydroxides of metals may be studied with the aid of the following activity series.

K Na Ca Ba	hydroxides of these metals are soluble in water, forming hydroxide solution (alkali); potassium and sodium hydroxides do not decompose on heating; calcium and barium hydroxides decompose on strong heating
Mg Al Zn Fe Sn Pb Cu	hydroxides of these metals are insoluble in water; on heating, they decompose into oxides and water
Hg Ag Au	hydroxides of these metals do not exist

1. Action of heat

Hydroxides of these metals in the above series (except potassium and sodium) decompose to oxide and water on heating.

$$\text{copper (II) hydroxide} \xrightarrow{\Delta} \text{copper (II) oxide} + \text{water}$$

$$Cu(OH)_2 \xrightarrow{\Delta} CuO + H_2O$$

2. Action with metals

Alkali solutions react with zinc and aluminium metals to give hydrogen and the zincates and aluminates.

zinc	+	sodium hydroxide	+	water	$\longrightarrow$	sodium zincate	+	hydrogen
Zn	+	$2NaOH$	+	$2H_2O$	$\longrightarrow$	$Na_2Zn(OH)_4$ (or) $Na_2ZnO_2.2H_2O$	+	H_2
aluminium	+	sodium hydroxide	+	water	$\longrightarrow$	sodium aluminate	+	hydrogen
$2Al$	+	$2NaOH$	+	$6H_2O$	$\longrightarrow$	$2NaAlO_2.2H_2O$ (or) $2NaAl(OH)_4$	+	$3H_2$

3. Action with acids

When hydroxides of metals react with acids, salts and water are formed.

calcium hydroxide	+	nitric acid	$\longrightarrow$	calcium nitrate	+	water
$Ca(OH)_2$	+	$2HNO_3$	$\longrightarrow$	$Ca(NO_3)_2$	+	$2H_2O$

SUMMARY

Oxygen combines directly with all other elements (except the noble gases and the noble metals) to form oxides. The main types of oxides are basic oxides (metallic oxides), acidic oxides (non- metallic oxides), amphoteric oxides (possess both acidic and basic properties), neutral oxides (neither acidic nor basic oxide), peroxides (react with acid to give salt and H_2O_2) and compound oxides (combination of two different oxides of the same element). Preparative methods for obtaining oxides from metals or their soluble salt are summarized in 11.3 on page 155. Metal oxides that are soluble in water react to produce the corresponding hydroxides. Hydroxides can be classified as basic hydroxides, amphoteric hydroxides and acidic hydroxides.

Questions and Problems

1. What are the main classes of oxides? Give one example of each class.
2. State to which class of oxide each of the following belongs.

   (a) copper (II) oxide
   (b) sulphur dioxide
   (c) aluminium oxide
   (d) carbon monoxide
   (e) di lead (II) lead (IV) oxide
   (f) barium peroxide

3. Describe the reactions, if any, of each of the following oxides with

(a) water (b) dilute nitric acid (c) sodium hydroxide solution.

(i) calcium oxide (ii) carbon monoxide (iii) zinc oxide

4. Describe how you would prepare :

(a) copper (II) oxide from copper (II) sulphate solution.

(b) zinc oxide from solid zinc carbonate.

(c) lead (II) oxide from lead (II) nitrate crystals.

5. The observations given below relate to three oxides, A, B and C. Classify the oxides, and in each case name one oxide which has the properties indicated.

A is a white crystalline solid which reacts vigorously with water, forming a solution which turns blue litmus paper red.

B is a white powder which is insoluble in water. It forms colourless solutions when separate portions are warmed with (i) dilute hydrochloric acid (ii) concentrated sodium hydroxide solution.

C is a white solid which reacts vigorously with water forming a white suspension. When this is filtered the filtrate turns red litmus paper blue.

6. Write TRUE or FALSE for each of the following statements.

(a) There are six main kinds of oxides.

(b) A basic oxide is a non-metallic oxide.

(c) An acidic oxide is a metallic oxide.

(d) Soluble acidic oxides dissolve in water to form acid solutions.

(e) An amphoteric oxide is a metallic oxide which possesses both basic and acidic properties.

(f) A neutral oxide which shows neither basic nor acidic character.

(g) Metallic oxide reacts with dilute acid to produce salt and hydrogen.

(h) A basic oxide is a metallic oxide.

7. Fill in the blanks with a suitable word or words.

(a) A compound oxide is an oxide, formed by the combination of two .... oxides of the same element.

(b) Metals or non-metals react with .... to give oxides.

(c) Some metallic hydroxides decompose to ....and water, on heating.

(d) Although water is neutral to litmus, it is not a ....oxide.

(e) Most of the carbonates decompose to oxides and ....on heating.

(f) An .... oxide is an oxide which possesses both acidic and basic properties.

(g) An .... oxide is a non-metallic oxide.

(h) Soluble basic oxides react with water , forming .... solutions.

8. Select the correct word or words given in the brackets.

(a) Some metals react with water to give (soluble, insoluble, different) hydroxides of metals and hydrogen.

(b) Some metals react with (hydrogen, oxygen, water) to give soluble hydroxides of metals and hydrogen.

(c) Addition of soluble alkalis (NaOH, KOH) to soluble salts of heavy metals give (soluble, insoluble, different) hydroxides.

(d) An acidic oxide is a (metallic, non-metallic, neutral) oxide.

(e) Soluble basic oxides react with water forming hydroxide solutions. These solutions turn (red, orange, violet) litmus blue.

(f) An alkali is a (basic, acidic, neutral) hydroxides which is soluble in water .

9. Match each of the items given in list A with the appropriate item given in list B.

List A	List B
(a) An acidic oxide	(i) which possesses both acidic and basic properties
(b) A basic oxide	(ii) an oxide which shows neither basic nor acidic character
(c) An amphoteric oxide	(iii) an oxide formed by the combination of two different oxides of the same element
(d) A compound oxide	(iv) a non-metallic oxide
(e) A neutral oxide	(v) a metallic oxide

10. Answer the following questions.

(a) What class of oxide does sulphur dioxide belong to?

Is it soluble in water? If so, what effect does it have on litmus paper?

(b) Explain the properties of basic oxides using the activity series.

****

CHAPTER 12

HYDROGEN

12.1 Occurrence

Only a small amount of hydrogen gas occurs in the atmosphere. In the combined state, however, hydrogen ranks as one of the more abundant elements. Water, acids and many organic substances reveal the wide- spread distribution of hydrogen in nature.

12.2 Preparation of Hydrogen

Laboratory preparation of hydrogen

Hydrogen is prepared in the laboratory by the action of the dilute mineral acids on certain metals.

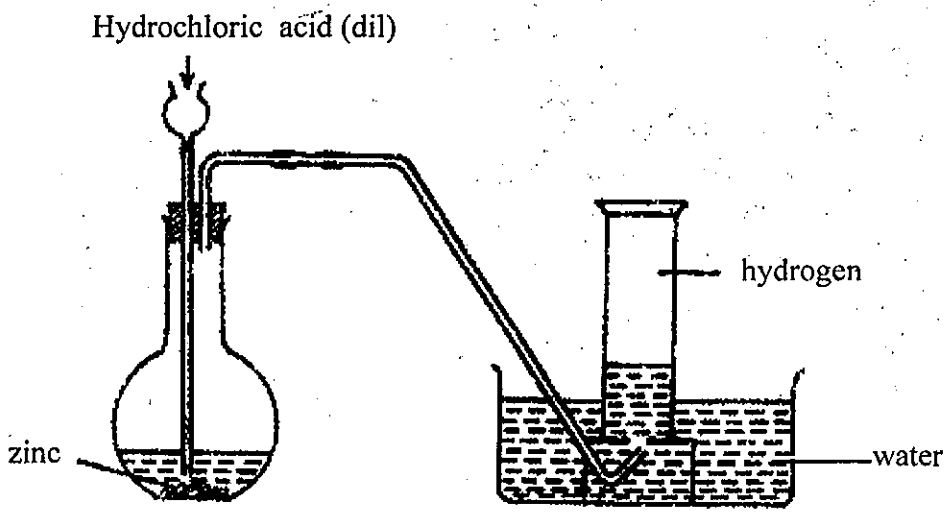


Fig. 12.1 Laboratory preparation of hydrogen

zinc + hydrochloric acid (dil) ⟶ zinc chloride + hydrogen

$$Zn + 2HCl \longrightarrow ZnCl_2 + H_2$$

Some granulated zinc is put into a flat-bottomed flask. Dilute hydrochloric acid is added to the zinc by means of a thistle funnel (Fig. 12.1).

Hydrogen gas is evolved and collected over water. Since hydrogen gas is slightly soluble in water, it is collected by the downward displacement of water.

Note (a) Dilute sulphuric acid may be used instead of dilute hydrochloric acid. The extent to which a given metal will react with a dilute acid of water to produce hydrogen depends on the position of the metal in the activity series as follows.

action with water		action with dilute acids
These metals react with cold water to produce hydroxide and hydrogen	K Na Ca	These metals (except Al) react with dilute sulphuric and hydrochloric acids to liberate hydrogen gas (Al reacts with hot concentrated acid only)
These metals react only with steam while heated, to produce oxide and hydrogen	Mg Al Zn Fe	
These metals react neither with water nor with steam	Pb (H) Cu Hg Ag Au	

(b) Dilute nitric acid is not used for the preparation of hydrogen by reaction with a metal. It will produce hydrogen, only in the case of reaction with magnesium.

(c) Sodium and potassium are never used with an acid to prepare hydrogen as the reactions are extremely violent.

(d) In practice, hydrogen gas initially coming out from a delivery tube, should not be collected because it is mixed with air. When it is mixed with air or oxygen and a naked flame or a light is applied, the mixture explodes.

Industrial preparation of hydrogen

There are two chief methods of manufacture of hydrogen.

(a) From hydrocarbons

Methane (natural gas) can be passed with steam over a nickel catalyst at 800°C and 30 atm pressure.

$$\text{methane} + \text{steam} \longrightarrow \text{carbon monoxide} + \text{hydrogen}$$

$$CH_4 + H_2O \xrightarrow[\text{800°C, 30 atm pressure}]{\text{Ni}} CO + 3H_2$$

The product is mixed with more steam and passed over iron (III) oxide catalyst at 450°C. Carbon monoxide is converted to carbon dioxide with further yield of hydrogen.

carbon monoxide + steam $\longrightarrow$ carbon dioxide + hydrogen

$$CO + H_2O \xrightarrow[450^\circ C]{Fe_2O_3} CO_2 + H_2$$

Carbon dioxide is removed by passing the gas mixture into water under 30 atm pressure (carbon dioxide gas is very soluble in water under pressure).

water + carbon dioxide $\overset{30\ atm}{\rightleftharpoons}$ carbonic acid

$$H_2O + CO_2 \rightleftharpoons H_2CO_3$$

Any remaining traces of carbon monoxide is removed by washing with a solution of copper (I) methanoate in ammonia.

(b) By electrolysis

Hydrogen is obtained as by-product in the electrolytic manufacture of chlorine from common salt or brine solution.

Hydrogen can also be made by electrolysis of water containing sulphuric acid.

water $\longrightarrow$ hydrogen + oxygen

$$2H_2O \longrightarrow 2H_2 + O_2$$

Test for hydrogen

Pure hydrogen burns quietly with a pale blue flame, but a mixture of hydrogen and air explodes when a flame is applied.

Note Since an air-hydrogen mixture explosion can be very dangerous, the test for hydrogen should always be carried out with very small volumes of the gas, i.e., the gas should be ignited in a test tube.

12.3 Properties of Hydrogen

1. Hydrogen is a colourless, odourless and tasteless gas. It is slightly soluble in water.

2. Hydrogen is much less dense than air.

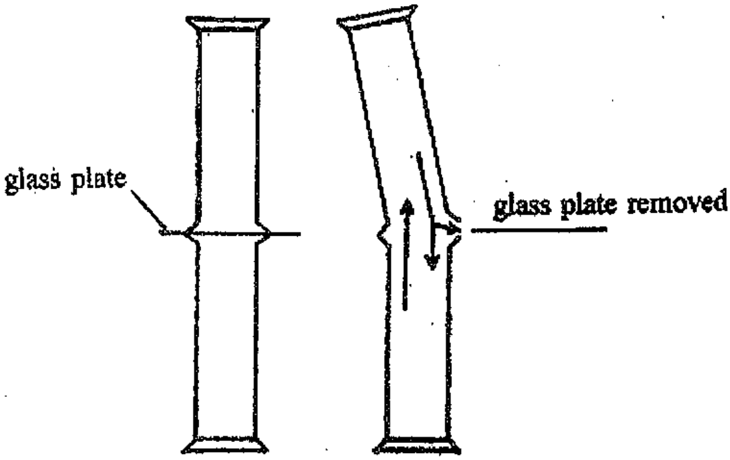


Fig. 12.2 Hydrogen is much less dense than air

A gas jar full of hydrogen is placed under a gas jar full of air. After a short time, the upper gas jar is removed (Fig. 12.2). Both gas jars are tested for hydrogen by applying a lighted splint. The contents of the upper gas jar will explode, showing that hydrogen has passed upward, displacing the air to form a mixture. The bottom gas jar will produce no explosion or burning, showing that the hydrogen previously in that jar has all been displaced by the air.

3. Hydrogen will burn, but it will not allow things to burn in it. (i.e., it will not support combustion.)

4. When hydrogen burns in air, water is formed.

$$\text{hydrogen} + \text{oxygen} \longrightarrow \text{water}$$

$$2H_2 + O_2 \longrightarrow 2H_2O$$

5. Hydrogen combines with most non-metals.

$$\text{hydrogen} + \text{sulphur} \longrightarrow \text{hydrogen sulphide}$$

$$H_2 + S \longrightarrow H_2S$$

$$\text{hydrogen} + \text{nitrogen} \longrightarrow \text{ammonia}$$

$$3H_2 + N_2 \xrightarrow[450^{\circ}C,\ 200\ atm]{Fe} 2NH_3$$

$$\text{hydrogen} + \text{chlorine} \longrightarrow \text{hydrogen chloride}$$

$$H_2 + Cl_2 \longrightarrow 2HCl$$

6. When hydrogen gas passes over heated oxides of metals (Fe and metals below it in the activity series), metal and water are formed.

copper (II) oxide (black)	+	hydrogen	$\xrightarrow{\Delta}$	copper (red)	+	water
CuO	+	H_2	$\xrightarrow{\Delta}$	Cu	+	H_2O
magnetic iron oxide	+	hydrogen	$\xrightarrow{\Delta}$	iron	+	water
Fe_3O_4	+	$4H_2$	$\xrightarrow{\Delta}$	$3Fe$	+	$4H_2O$

Note The type of reaction just described is known as reduction of the metallic oxide to metal.

Reduction may be defined as the removal of oxygen from a substance or the addition of hydrogen to it.

A substance, which brings about the reduction, is a reducing agent.

12.4 Uses of Hydrogen

Since hydrogen is the lightest gas known, it has been used to inflate balloons, like weather balloons which float in the air. It is also used in the conversion of oils to solid fats, in the conversion of coal to synthetic petrol, in the manufacture of ammonia by Haber's Process and in the synthesis of hydrochloric acid. It is used as an oxyhydrogen torch in the cutting and welding of steel.

SUMMARY

Hydrogen is a colorless, odourless, tasteless gas which is present only a small amount in atmosphere. It is prepared in the laboratory by the action of the dilute mineral acids (except HNO_3) on certain metals. With the exception of K and Na, liberation of hydrogen depends on the position of metals in the activity series. Dilute HNO_3 can produce hydrogen gas only in the case of reaction with Mg. There are two chief methods of manufacture of hydrogen: (i) from hydrocarbon (ii) by-product in the electrolytic manufacture of chlorine from common salt. When hydrogen burns in air, water is formed. It combines with most non-metals and reduces the metallic oxide to metal. The usage of hydrogen are weather ballons; the conversion of oils to solid fats, and coal to synthetic petrol.

Questions and Problem

1. Draw a fully labelled diagram to illustrate the laboratory preparation and collection of hydrogen by the action of metal on an acid.

2. Describe the reactions, if any, of each of the following metals with water or steam.

   (a) zinc (b) copper (c) sodium

   On the basis of these reactions, arrange the above three metals in order of increasing chemical reactivity.

3. With what results and under what conditions does hydrogen react with

   (a) nitrogen (b) sulphur (c) chlorine?

4. How would you prepare hydrogen gas in the laboratory by using magnesium metal?

5. (a) Write equations for the reactions of dilute acids with

   (i) K (ii) Na (iii) Ca

   (b) Explain why the reactions of these metals and dilute acids are not used in preparing hydrogen.

   (c) Why do we not use nitric acid in preparing hydrogen?

6. Write TRUE or FALSE for each of the following statements.

   (a) A small amount of hydrogen occurs in the atmosphere.

   (b) Hydrogen is prepared in the laboratory by the action of dilute mineral acids on certain non-metals.

   (c) Since hydrogen gas is insoluble in water, it is collected by the downward displacement of water.

   (d) The metals such as Pb, Cu, Hg, Au, Ag react with water or steam.

   (e) Na, K are never used with an acid to produce hydrogen.

7. Fill in the blanks with a suitable word or words.

   (a) Hydrogen gas is prepared by the action of dilute hydrochloric acid on ....

   (b) Hydrogen gas is manufactured by passing .... (natural gas) with steam over a nickel catalyst at 800°C and 30 atm pressure.

   (c) Hydrogen is obtained as by-product in the electrolytic manufacture of .... from common salt or brine solution.

   (d) Hydrogen is a colourless, odourless and .... gas.

(e) Hydrogen is much less .... than air.

(f) Hydrogen can be prepared by the action of ....on heated iron.

(g) An atom of hydrogen is the .... of all atoms.

8. Select the correct word or words given in the brackets.

(a) A substance which brings about the reduction, is a/an (oxidizing, reducing, decomposing) reagent.

(b) Hydrogen gas is evolved and collected (above, under, over) water.

(c) Since hydrogen gas is insoluble in water, it is collected by (upward, downward, under) displacement of water.

(d) The metals (K, Ca, Na) react with cold water to produce hydroxide and (oxygen, carbon dioxide, hydrogen).

(e) When hydrogen burns in air, (nitrogen, carbon dioxide, water) is formed.

(f) When hydrogen burns in air (alkali, acid, water) is formed.

(g) Which of the following combinations react to produce hydrogen ?

(i) copper and hydrochloric acid

(ii) copper oxide and sulphuric acid

(iii) calcium and water

(h) Hydrogen gas can be used as a fuel. When it burns, it

(i) produces a smoky flame.

(ii) produce water and carbon dioxide.

(iii) explodes violently.

(iv) produces water only.

9. Answer the following questions.

(a) How will you test for hydrogen gas ?

(b) Write down any two uses of hydrogen.

10. Match each of the items given in **list A** with the appropriate item given in ~~**list B.**~~

List A	**List B**
(a) The metals (K, Na, Ca) are reacted with cold water	(i) hydrogen is obtained as by-product
(b) The dilute mineral acid is reacted with certain metals	(ii) it has been used to inflate balloons, like weather balloons which float in the air
(c) Hydrogen is the lightest gas known	(iii) hydroxide and hydrogen are produced
(d) The electrolytic manufacture of chlorine from common salt or brine solution	(iv) hydrogen gas is evolved

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CHAPTER 13

WATER

13.1 Importance of Water

Water plays a vital role in various physical and chemical processes upon which life depends. Life is dependent on the storing of energy from the sun in sugars formed by the following reaction, called photosynthesis.

$$\text{carbon dioxide} \quad + \quad \text{water} \xrightarrow[\text{light energy}]{h\nu} \text{sugar} \quad + \quad \text{oxygen}$$

$$6CO_2 \quad + \quad 6H_2O \longrightarrow C_6H_{12}O_6 \quad + \quad 6O_2$$

This reaction goes on in all green plants in the presence of sunlight. The energy stored in plants or animal bodies is released in the reaction called respiration which in a sense is the reverse of photosynthesis.

$$\text{sugar} \quad + \text{oxygen} \longrightarrow \text{carbon dioxide} + \text{water} + \text{energy}$$

$$C_6H_{12}O_6 \quad + \quad 6O_2 \longrightarrow 6CO_2 \quad + 6H_2O + \text{energy}$$

Water is a major constituent of our bodies. Our bodies contain about 70% of water. Tomatoes contain over 90% of water and potatoes, corn, eggs and beef contain about 75% of water.

The processes by which nutrient materials are taken up by plants and animals, (the metabolic processes in plants and animals) and the elimination of waste products by animal organisms take place largely in aqueous solution or suspension.

In addition, water is an important industrial raw material.

13.2 Occurrence of Water

Water is the most abundant and most widespread compound on earth. It occurs as water vapour in the atmosphere. In the liquid state water forms the oceans, rivers and lakes that cover three quarters of the earth. Water also occurs as ground water in the interstices of the soil and rocks.

13. Natural Waters

Pure water does not exist in a neutral state. All natural waters contain impuri s in varying amounts.

Rain water

Rain water is the purest form of natural water. However, it may contain dissolved gases (such as oxygen, nitrogen and carbon dioxide) and dust from the atmosphere. Dissolved carbon dioxide makes rain water slightly acidic. The rain water may even contain dissolved harmful gases, such as sulphur dioxide and oxides of nitrogen, as it falls through the air over industrial areas.

River Water

River water contains some dissolved gases and also some dissolved solids depending upon the soil over which it passes. River water is often unfit to drink, particularly if it has passed by a town pouring sewage into it, or by factories dumping waste toxic chemicals.

Sea water

The river finally flows into the sea carrying contamination with it. Sea water, therefore, is the reservoir into which all the impurities are collected. Sea water contains various salts dissolved in it.

13.4 Physical Properties of Water

1. Pure water is a clear, transparent liquid, colourless in thin layers. Thick layers of water have bluish colour.

2. The freezing point of water is 0° C, and the boiling point of water at 1 atmosphere is 100° C.

3. With decrease in temperature, most substances diminish in volume, and hence increase in density. Water has the very unusual property of having a temperature at which its density is a maximum. This temperature is 3.98°C or 4°C to the nearest degree. (See Table 13.1)

Table 13.1 Density of water at various temperatures

Temperature (°C)	Density (g cm^{-3})
0.00	0.99987
3.98	1.00000
10.00	0.99973
20.00	0.99823
100.00	0.95838

4. Water is regarded as the most universal solvent. The reason why it is vary difficult to prepare perfectly pure water is due to the fact that water dissolves almost

all substances to a greater or lesser extent. Because of this property, water is known as the universal solvent.

13.5 Chemical Properties of Water

1. Action of metals with water

K, Na	violent reaction with water
Ca	slow reaction with water
Mg, Al, Zn, Fe	react with steam
Pb, (H), Cu, Hg, Ag, Au	do not react with water or steam

potassium + water $\longrightarrow$ potassium hydroxide + hydrogen

$$2K + 2H_2O \longrightarrow 2KOH + H_2$$

sodium + water $\longrightarrow$ sodium hydroxide + hydrogen

$$2Na + 2H_2O \longrightarrow 2NaOH + H_2$$

calcium + water $\longrightarrow$ calcium hydroxide + hydrogen

$$Ca + 2H_2O \longrightarrow Ca(OH)_2 + H_2$$

Zinc does not react with hot or cold water. When a current of steam is passed over red hot zinc, zinc oxide and hydrogen are formed.

zinc + steam $\longrightarrow$ zinc oxide + hydrogen

$$Zn + H_2O \longrightarrow ZnO + H_2$$

At ordinary temperature, iron does not react with water. Rust is formed only in the presence of air. Iron reacts with an excess of steam at red heat to form tri iron tetroxide and hydrogen.

$$\text{iron} + \text{steam} \longrightarrow \text{tri iron tetroxide} + \text{hydrogen}$$

$$3Fe + 4H_2O \longrightarrow Fe_3O_4 + 4H_2$$

2. Action of non-metal with water

Carbon reacts with steam at white heat to form carbon monoxide and hydrogen. This mixture of carbon monoxide and hydrogen is known as water gas.

$$\text{carbon} + \text{steam} \longrightarrow \text{water gas}$$

$$C + H_2O \longrightarrow (CO+H_2)$$

Chlorine reacts with water forming a mixture of two acids.

$$\text{chlorine} + \text{water} \longrightarrow \text{hypochlorous acid} + \text{hydrochloric acid}$$

$$Cl_2 + H_2O \longrightarrow HOCl + HCl$$

3. Action of water on oxides

Metallic oxides such as oxides of potassium, sodium, calcium react with water forming corresponding hydroxides.

$$\text{potassium oxide} + \text{water} \longrightarrow \text{Potassium hydroxide}$$

$$K_2O + H_2O \longrightarrow 2KOH$$

$$\text{sodium oxide} + \text{water} \longrightarrow \text{sodium hydroxide}$$

$$Na_2O + H_2O \longrightarrow 2NaOH$$

Calcium oxide (quicklime) also reacts with water to form calcium hydroxide which is also known as slaked lime. Calcium hydroxide is sparingly soluble in water.

$$\underset{\text{(quicklime)}}{\text{calcium oxide}} + \text{water} \longrightarrow \underset{\text{(slaked lime)}}{\text{calcium hydroxide}}$$

$$CaO + H_2O \longrightarrow Ca(OH)_2$$

Non-metallic oxides such as sulphur dioxide, sulphur trioxide, nitrogen dioxide, NO_2 and carbon dioxide react with water on dissolving in it, to form the corresponding acids.

$$\text{sulphur dioxide} + \text{water} \longrightarrow \text{sulphurous acid}$$

$$SO_2 + H_2O \longrightarrow H_2SO_3$$

sulphur trioxide	+	water	$\longrightarrow$	sulphuric acid
SO_3	+	H_2O	$\longrightarrow$	H_2SO_4
carbon dioxide	+	water	$\longrightarrow$	carbonic acid
CO_2	+	H_2O	$\longrightarrow$	H_2CO_3

4. Action of water on metallic chlorides

Some metallic chlorides, eg., magnesium, aluminium, zinc and iron chlorides are hydrolysed by water. When the solutions of these chlorides are evaporated, metal chlorides are not recovered. A basic salt or the hydroxide or oxide of the metal is formed. Anhydrous aluminium chloride reacts violently with water forming the hydrate, $AlCl_3.6H_2O$. In excess of water, it dissolves to form aluminium hydroxide. This is another example of the hydrolysis of salts.

aluminium chloride	+	water	$\longrightarrow$	aluminium hydroxide	+	hydrochloric acid
$AlCl_3$	+	$3H_2O$	$\longrightarrow$	$Al(OH)_3$	+	$3HCl$

When this solution is evaporated to dryness, the residue is aluminium hydroxide, not aluminium chloride.

When iron (III) chloride is dissolved in water, iron (III) hydroxide is formed.

On evaporation of this solution, the residue will be iron (III) oxide.

iron (III) chloride	+	water	$\longrightarrow$	iron (III) hydroxide	+	hydrogen chloride
$FeCl_3$	+	$3H_2O$	$\longrightarrow$	$Fe(OH)_3$	+	$3HCl$

iron (III) hydroxide	$\xrightarrow{\Delta}$	iron (III) oxide	+	steam
$2Fe(OH)_3$	$\xrightarrow{\Delta}$	Fe_2O_3	+	$3H_2O$

5. Action of water on metallic carbides

Some metallic carbides such as calcium and aluminium carbides react with water to form metal hydroxide and hydrocarbon.

calcium carbide	+	water	$\longrightarrow$	calcium hydroxide	+	ethyne
CaC_2	+	$2H_2O$	$\longrightarrow$	$Ca(OH)_2$	+	C_2H_2
aluminium carbide	+	water	$\longrightarrow$	aluminium hydroxide	+	methane
Al_4C_3	+	$12H_2O$	$\longrightarrow$	$4Al(OH)_3$	+	$3CH_4$

13.6 Hardness of Water

Natural waters are never pure. Rain water contains dissolved gases and dust from the atmosphere and starts to dissolve other substances as soon as it reaches the ground. In fresh water the presence of certain kinds of dissolved salt is referred to as hardness.

Water is said to be hard if it contains a significant concentration of salts that react with soap to produce a curdy precipitate before a permanent lather is formed.

Hard water is water which will not readily form a permanent lather with soap.

Soft water is water which readily gives a permanent lather with soap.

Many salts dissolved in water can cause hardness, but the ones most common are salts of calcium and magnesium (and sometimes iron (II)). The production of curdy precipitate is wasteful of soap. The amount of soap that is wasted can be considerable in commercial laundry. Furthermore, dissolved salt of calcium, magnesium and iron will precipitate in kettles and boilers of industrial steam generating equipment to produce a hard scale. Scale causes wastage of heat and localized heating, which in extreme cases, may result in explosion. Hard water can also cause interferences in processes such as dyeing and tanning. Iron compounds stain white fabrics pottery, and enamel ware. For the above reasons the hardness of water is always objectionable.

Soap

Soap is the sodium salt of organic acid. The most common one is sodium stearate which is the sodium salt of stearic acid, $C_{17}H_{35}COOH$. The formula of sodium stearate is $C_{17}H_{35}COONa$. Formula of sodium stearate is also represented as Na(St). The St is used for the group $C_{17}H_{35}COO^-$.

If water contains calcium salts such as calcium hydrogen carbonate, dissolved in it, soap is precipitated as calcium stearate which appears as a curd.

sodium stearate (soap)	+	calcium hydro-gencarbonate	$\longrightarrow$	sodium hydro gencarbonate	+	calcium stearate
$2Na(St)$	+	$Ca(HCO_3)_2$	$\longrightarrow$	$2NaHCO_3$	+	$Ca(St)_2$

Detergent

Soap is not the only, cleaning agent used. So called soapless soaps (or detergents) are available and are used extensively in everyday life and also in industry. Soap is made from fats and oils, but detergents are not made from fats and oils. These soapless detergents contain molecules with a salt-like group attached to a long chain of hydrocarbon. (Hydrocarbons are the compounds which contain carbon and hydrogen only.) The advantage of these detergents is that their salts of calcium and magnesium are soluble in water. Therefore, when used in hard water detergents do not form curdy or greasy scum.

There are two types of hardness depending upon the types of salts dissolved in water.If the hardness of water can be removed by boiling it is said to be temporarily hard, but if it cannot be removed by boiling it is said to be permanently hard.

Temporary hardness and its removal

Water is said to have temporary hardness when dissolved material consists mainly of calcium hydrogencarbonate, or magnesium hydrogencarbonate or both. Boiling such water causes the precipitation of normal carbonates of calcium and magnesium.

calclum hydrogencarbonate	$\xrightarrow{\Delta}$	calcium carbonate	+	water	+	carbon dioxide
$Ca(HCO_3)_2$	$\xrightarrow{\Delta}$	$CaCO_3\downarrow$	+	H_2O	+	CO_2

magnesium hydrogencarbonate	$\xrightarrow{\Delta}$	magnesium carbonate	+	water	+	carbon dioxide
$Mg(HCO_3)_2$	$\xrightarrow{\Delta}$	$MgCO_3\downarrow$	+	H_2O	+	CO_2

The carbonates can be filtered off, and the filtered water which no longer contains the cause of hardness (soluble hydrogencarbonate of calcium and magnesium) is said to be softened. The removal of temporary hardness by boiling is too expensive to be done on a large scale. A cheaper method involves adding a calculated quantity of slaked lime (excess of lime will cause hardness).

slaked lime	+	calcium hydrogencarbonate	$\longrightarrow$	calcium carbonate	+	water
$Ca(OH)_2$	+	$Ca(HCO_3)_2$	$\longrightarrow$	$2CaCO_3$	+	$2H_2O$
(slightly soluble)				(insoluble)		

The precipitated calcium carbonate may be filtered off. This process is known as the **Clark's Process.**

Permanent hardness and its removal

If calcium and magnesium are not present as the hydrogencarbonate, but as the chloride or sulphate, then boiling will not remove the hardness. This type of water is said to have permanent hardness. It can be softened by appropriate treatment.

The methods of softening involve the addition of chemicals that either precipitate the calcium and magnesium ions or convert them into complex ions that will not precipitate with soap.

One of the most common precipitating agents is sodium carbonate (washing soda, $Na_2CO_3 . 10H_2O$). Sodium carbonate reacts with calcium and magnesium compound according to the following reactions, with the precipitation of insoluble carbonates.

$$\text{calcium chloride} + \text{sodium carbonate} \longrightarrow \text{calcium carbonate} + \text{sodium chloride}$$

$$CaCl_2 + Na_2CO_3 \longrightarrow CaCO_3 \downarrow + 2NaCl$$

$$\text{magnesium sulphate} + \text{sodium carbonate} \longrightarrow \text{magnesium carbonate} + \text{sodium sulphate}$$

$$MgSO_4 + Na_2CO_3 \longrightarrow MgCO_3 \downarrow + Na_2SO_4$$

After filtration of the precipitated carbonates, the hard water becomes soft.

13.7 Purification of Water

Purification by distillation

Ordinary water is more or less impure; it usually contains dissolved salts and dissolved gases and sometimes organic matter. For medicinal and laboratory use water is purified by distillation. Pure tin vessels and pipes are often used for storing and transporting distilled water.

Removal of ionic impurities from water (Deionization of water)

Ionic impurities can be effectively and cheaply removed from water by passing through substances like zeolite which have giant structure. A zeolite is an aluminosilicate (or sodium aluminium silicate), with formula such as $NaAl_2Si_4O_{12}$. It consists of a rigid framework formed by the aluminium, silicon and oxygen atoms. But sodium ions are loosely held and may be replaced. When hard water flows over zeolite grains, some of the sodium ions are replaced by ions of calciunm, magnesium and iron which are present in hard water. The removal of calcium, magnesium and iron results in removal of hardness. After most of the sodium ions have been replaced, the spent zeolite is regenerated by allowing it to stand in contact with a saturated sodium chloride solution (brine); the reaction is reversed, Na^+ replacing Ca^{2+} and the other cations.

The reaction that occurs may be written using symbols.

sodium zeolite + calcium chloride ⟶ calcium zeolite + sodium chloride

$$2NaZ + CaCl_2 \longrightarrow CaZ_2 + 2NaCl$$

When concentrated sodium chloride solution (brine) is run through the spent zeolite, the reverse reaction occurs.

sodium chloride + calcium zeolite ⟶ sodium zeolite + calcium chloride

$$2NaCl + CaZ_2 \longrightarrow 2NaZ + CaCl_2$$

The advantage of using substances like zeolite which have giant structure containing the aluminosilicate framework is that these substances, are not carried along in the water. They remain in the water softening tank.

This method of softening water is also known as the **Permutit Method.** Permutit is a complex sodium aluminium silicate. It is also known as zeolite when found as a mineral.

SUMMARY

In this chapter, there are seven main topics concerning with water to be studied:such as the importance of water, occurrence of water, physical and chemical properties of water, hardness of water, including some aspects of soap and detergent, and finally purification of water.

Questions and Problems

1. Write equations to show five different elements reacting with water. Mention the conditions.
2. Write equations to show five different compounds reacting with water. Mention the conditions.
3. How will you test for hardness in water?

   Explain how you will distinguish between temporary and permanent hardness in water.
4. Which one of the following do you consider to be "hard" ?
   (a) Potassium chloride solution
   (b) Distilled water
   (c) Rain water with dissolved air
   (d) Sea water

5. (a) Name two compounds in each case which can cause

(i) temporary hardness
(ii) permanent hardness.

(b) Show, by equations, how both temporary and permanent hard water react with soap.

6. Describe three methods which can be used for softening hard water.

7. In the electrolysis of water, 35.6 cm^3 of oxygen is collected at the anode. What volume of hydrogen will be collected at the cathode? (Assume that the solubility of each of the two gases in water is very small.)

8. How many grams of triiron tetroxide will be obtained when steam reacts with 100 g of heated iron?

9. What mass of hydrogen is combined with oxygen in 4.2 g of water?

10. If 224 dm^3 of hydrogen measured at STP are exploded in excess oxygen. What is the mass of water produced?

11. Write TRUE or FALSE for each of the following statements.

(a) Water is a major constituent of our bodies.
(b) Sea water contains various salts dissolved in it.
(c) Water is not regarded as the most universal solvent.
(d) Water react violently with K, Na metals.
(e) Our bodies contain about 75% of water.
(f) Silicon dioxide is soluble in water.
(g) Metals such as Pb, Cu, Hg, Au, Ag react with water or steam.
(h) Pure water can exist in a neutral state.

12. Fill in the blanks with a suitable word or words.

(a) Carbon reacts with steam at white heat to form ....and hydrogen.

(b) Calcium oxide (quick lime) also reacts with water to form hydroxide which is also known as ....

(c) Hard water is water, which will not readily form a permanent lather with ....

(d) Soap is the .... of organic acid.

(e) If the hardness of water can be removed by ...., it is said to be temporarily hard.

13. Select the correct word or words given in the brackets.

(a) Water is said to have temporary hardness when (dissolved, insoluble, some) material consists mainly of calcium and magnesium hydrogen carbonate.

(b) Sulphur dioxide gas reacts with water on dissolving in it to form (sulphurous, sulphuric, carbonic) acid.

(c) Soft water is water which readily gives a permanent (precipitate, lather, material) with soap.

(d) Calcium oxide reacts with water to form calcium hydroxide which is known as (hydrated lime, slaked lime, quick lime).

(e) The reaction goes on in all green plants in the presence of sunlight is known as (respiration, photosynthesis, photodehydration).

(f) Some metals react with water to give (soluble, insoluble, different) hydroxide of metals and hydrogen.

(g) (Sea, River, Rain) water is hard water.

14. Answer the following questions.

(a) What are some of the uses of water in industry and in home?

(b) Approximately how much water do you drink in a day?

(c) Explain the action of metals with water based on activity series.

15. Match each of the items given in list A with the appropriate item given in list B.

List A	List B
(a) Potassium reacts violently with water	(i) temporary hardness and permanent hardness.
(b) The sodium salt of organic acid	(ii) to form carbon monoxide and hydrogen
(c) Two types of hardness	(iii) to form potassium hydroxide and hydrogen
(d) Carbon reacts with steam at white heat	(iv) soap
(e) Removal of ionic impurities from water	(v) deionization of water

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CHAPTER 14

ACIDS, BASES AND SALTS

14.1 Acids

When an acid is dissolved in water, it splits up into electrically charged particles called ions. This process is known as dissociation.

$$\text{Hydrochloric acid } HCl \longrightarrow H^+ + Cl^-$$
$$\text{Sulphuric acid } H_2SO_4 \longrightarrow H^+ + HSO_4^-$$
$$HSO_4^- \longrightarrow H^+ + SO_4^{2-}$$
$$\text{Nitric acid } HNO_3 \longrightarrow H^+ + NO_3^-$$

Note that the dissociation of all acids give H^+ ion.

However, when an acid dissociates in water, it does not yield a simple hydrogen ion. Instead, it forms a hydrated hydrogen ion by a reaction between H_2O and H^+ ion.

$$H_2O \quad + \quad H^+ \rightleftharpoons H_3O^+$$

This hydrated hydrogen ion or hydronium is responsible for acidity. For ordinary purposes the hydration of the hydrogen ion is often ignored and H^+ is used to indicate the product of the dissociation. An acid, therefore, may be defined as follow:

An acid is a compound which, when
dissolved in water, produces hydrogen ions, H^+,
as the only positive ion.

14.2 Properties of Acids

1. Most dilute acids have a sour taste.
2. Most acids turn blue litmus red.
3. Metals which are much more electropositive than hydrogen (Chapter 12) react with dilute hydrochloric or dilute sulphuric acid to liberate hydrogen. Metals which are only slightly more electropositive than hydrogen, react with neither of these dilute acids, but will react with hot concentrated acid., Nitric acid is a strong oxidizing agent and will not produce hydrogen on reaction with metals (except Mg). Oxides of nitrogen are only obtained.

4. Almost all acids react with carbonates to liberate carbon dioxide.

5. All acids react with bases, forming salts and water.

14.3 Methods of Preparation of Acids

Acids may be prepared by the following methods

1. By the reaction between the acidic oxide of a non-metal and water

Example: Sulphurous acid and phosphoric acid can be prepared by the action of sulphur dioxide and phosphorus (V) oxide with cold water.

sulphur dioxide + water $\longrightarrow$ sulphurous acid

$$SO_2 + H_2O \longrightarrow H_2SO_3$$

phosphorus (V) oxide + water $\longrightarrow$ phosphoric acid

$$P_4O_{10} + 6H_2O \longrightarrow 4H_3PO_4$$

2. By displacing a weaker or more volatile acid from its salt by a stronger or less volatile acid

Example (a) : Displacement of the more volatile hydrogen chloride from metallic chloride by the less volatile concentrated sulphuric acid

sodium chloride + sulphuric acid $\longrightarrow$ sodium hydrogensulphate + hydrogen chloride

$$NaCl + H_2SO_4 \longrightarrow NaHSO_4 + HCl$$

Example (b) : Displacement of the weaker boric acid from borax by sulphuric acid

borax + sulphuric acid + water $\longrightarrow$ sodium sulphate + boric acid

$$Na_2B_4O_7 + H_2SO_4 + 5H_2O \longrightarrow Na_2SO_4 + 4H_3BO_3$$

3. By precipitating an insoluble sulphide from a metallic salt with hydrogen sulphide

Example:

lead (II) ethanoate + hydrogen sulphide $\longrightarrow$ lead (II) sulphide + ethanoic acid

$$(CH_3COO)_2Pb + H_2S \longrightarrow PbS + 2CH_3COOH$$

14.4 Bases

A base is usually a metallic oxide (or) hydroxide and will react with an acid to form a salt and water only.

Example :

copper (II) oxide	+	sulphuric acid	$\longrightarrow$	copper(II) sulphate	+	water
CuO	+	H_2SO_4	$\longrightarrow$	$CuSO_4$	+	H_2O
sodium hydroxide	+	hydrochloric acid	$\longrightarrow$	sodium chloride	+	water
$NaOH$	+	HCl	$\longrightarrow$	$NaCl$	+	H_2O

14.5 Alkalis

An alkali is a basic hydroxide which is soluble in water. The nature of the hydroxide of the metal varies according to the position of the metal in the activity series, as illustrated below.

K, Na, Ca	The hydroxides of these metals are soluble in water and are alkalis.	Hydroxides of sodium and potassium do not decompose on heating
Mg, Al, Zn, Fe, Pb	The hydroxides of these metals are insoluble in water. These hydroxides (except of Fe and Mg) are amphoteric.	Hydroxides of these metals decompose into oxide and water on heating
Cu	Hydroxide is insoluble in water	
Hg, Ag, Au	Hydroxides of these metals do not exist.	

14.6 Properties of Bases

1. Soluble or insoluble bases react with acids to form salts and water only.
2. When a basic oxide is soluble in water, it is known as an alkali. It has the following properties :
   - (a) It turns red litmus blue.
   - (b) It is soapy or slippery to the touch.
   - (c) It reacts with acid to form a salt and water.

14.7 Neutralization

Neutralization is the reaction between a basic oxide or hydroxide and an acid, resulting in the formation of a salt and water.

14.8 Salts

A salt is a compound, produced from the reaction of a base or metal with an acid. In other words, a salt is a compound consisting of a positively charged metallic ion and a negatively charged ion which is derived from the corresponding acid by loss of H^+.

The most common types are as follows :

(a) A normal salt is one which is made up of only a metallic radical, united with an acid radical. It contains neither replaceable hydrogen, nor a hydroxyl group.

(b) An acid salt is formed when only part of the hydrogen of an acid has been replaced by a metal, such salts still contain replaceable hydrogen.

Example :

Acid	Salt	Metallic or basic radical	Acid radical
H_2SO_4	Acid salt $NaHSO_4$	Na^+	HSO_4^-
	Normal salt Na_2SO_4	Na^+	$SO_4^=$
H_2CO_3	Acid salt $NaHCO_3$	Na^+	HCO_3^-
	Normal Salt Na_2CO_3	Na^+	$CO_3^=$
H_2S	Acid salt $NaHS$	Na^+	HS^-
	Normal salt Na_2S	Na^+	$S^=$

(c) A basic salt is one which contains one or more hydroxyl groups, besides having an acid radical.

Example :

Basic hydroxide	Basic salt	Normal salt
$Zn(OH)_2$	$Zn(OH)Cl$	$ZnCl_2$
$Mg(OH)_2$	$Mg(OH)Cl$	$MgCl_2$

14.9 Methods of Preparation of Salts

Salts are prepared in a variety of ways. The method chosen for preparing any particular salt depends largely on whether it is soluble in water or not. Soluble salts are usually prepared by methods which involve crystallization. Insoluble salts are usually prepared by methods which involve precipitation.

	Soluble salts		*Insoluble salts*
1.	All nitrates of metals		
2.	All chlorides (except those in the next column)	2.	Silver chloride; mercury (I) chloride; (lead (II) chloride is soluble only in hot water)
3.	All sulphate (except those in the next column)	3.	Barium sulphate and lead (II) sulphate. (Calcium sulphate is sparingly soluble)
4.	Sodium, potassium and ammonium carbonates	4.	All carbonates except those of sodium, potassium and ammonium

Method 1 Action of an acid on a metal

Soluble salts may be prepared by using a dilute acid and a metal. The salt formed then passes into solution in the aqueous acid and can be separated by crystallization.

Example:

zinc + sulphuric acid (dil) $\longrightarrow$ zinc sulphate + hydrogen

$$Zn + H_2SO_4 \longrightarrow ZnSO_4 + H_2$$

Zinc sulphate crystallizes out from aqueous solution as zinc sulphate heptahydrate, $ZnSO_4 \cdot 7H_2O_{(s)}$.

Some salts can be prepared by using a metal and a concentrated acid.

copper	+	sulphuric acid	$\longrightarrow$	copper (II) sulphate	+	water	+	sulphur dioxide
Cu	+	$2H_2SO_4$	$\longrightarrow$	$CuSO_4$	+	$2H_2O$	+	SO_2

Copper (II) sulphate crystallizes out from aqueous solution as copper (II) sulphate pentahydrate, $CuSO_4.5H_2O_{(s)}$.

Nitrates of certain metals can be prepared by the action of dilute or concentrated nitric acid upon metals. The nitrates of common heavy metals are very soluble in water. These nitrates, except lead (II) nitrate are deliquescent. This fact makes it difficult to prepare their crystals.

Example :

lead	+	nitric acid (dil)	$\longrightarrow$	lead (II) nitrate	+	water	+	nitrogen oxide
3Pb	+	$8HNO_3$	$\longrightarrow$	$3Pb(NO_3)_2$	+	$4H_2O$	+	2NO
copper	+	nitric acid	$\longrightarrow$	copper (II) nitrate	+	water	+	nitrogen dioxide
Cu	+	$4HNO_3$	$\longrightarrow$	$Cu(NO_3)_2$	+	$2H_2O$	+	$2NO_2$

Chlorides of heavy metals are generally prepared in the anhydrous state by heating the metal in a current of dry chlorine or hydrogen chloride. The reason for this is that many of them crystallize with water of crystallization. If an attempt is made to drive off this water by heating the crystalline salt, the water reacts with normal salt to form basic salt. This reaction is known as hydrolysis.

zinc chloride mono hydrate	$\xrightarrow{\Delta}$	zinc hydroxy chloride	+	hydrogen chloride
$ZnCl_2.H_2O$	$\xrightarrow{\Delta}$	$Zn(OH)Cl$	+	HCl

Note An anhydrous salt is a compound with no water of crystallization.

Example: Hydrated salt of copper (II) sulphate $CuSO_4.5H_2O$

Anhydrous salt of copper (II) sulphate $CuSO_4$

Method 2 By double decomposition

In a double decomposition reaction, two soluble compounds are used as starting materials to produce one soluble compound and one insoluble compound. The insoluble compound can be easily separated by filtering.

Example :

lead (II) nitrate	+	sodium sulphate	$\longrightarrow$	lead (II) sulphate	+	sodium nitrate
$Pb(NO_3)_2$	+	Na_2SO_4	$\longrightarrow$	$PbSO_4$	+	$2NaNO_3$
barium chloride	+	sulphuric acid (dil)	$\longrightarrow$	barium sulphate	+	hydrogen chloride
$BaCl_2$	+	H_2SO_4	$\longrightarrow$	$BaSO_4$	+	$2HCl$
calcium chloride	+	sodium carbonate	$\longrightarrow$	calcium carbonate	+	sodium chloride
$CaCl_2$	+	Na_2CO_3	$\longrightarrow$	$CaCO_3$	+	$2NaCl$

Method 3 By neutralization

Neutralization is the reaction between an acid and a base to produce a salt and water. The use of this process depends on the solubility of base in water.

Preparation of a salt from an alkali (soluble base)

Salts of sodium, potassium and ammonium can be prepared from caustic soda (sodium hydroxide), caustic potash (potassium hydroxide), and ammonia solution respectively by the neutralization reaction using the appropriate acid.

Example :

sodium hydroxide	+	hydrochloric acid (dil)	$\longrightarrow$	sodium chloride	+	water
$NaOH$	+	HCl	$\longrightarrow$	$NaCl$	+	H_2O
potassium hydroxide	+	sulphuric acid (dil)	$\longrightarrow$	potassium sulphate	+	water
$2KOH$	+	H_2SO_4	$\longrightarrow$	K_2SO_4	+	$2H_2O$
potassium hydroxide	+	nitric acid (dil)	$\longrightarrow$	potassium nitrate	+	water
KOH	+	HNO_3	$\longrightarrow$	KNO_3	+	H_2O

ammonium hydroxide	+	nitric acid (dil)	$\longrightarrow$	ammonium nitrate	+	water
NH_4OH	+	HNO_3	$\longrightarrow$	NH_4NO_3	+	H_2O

Preparation of a soluble salt from an insoluble base

If a base used in neutralization is insoluble in water (i.e., not an alkali), the above method cannot be applied. In such cases the salts of metals can be prepared by using either the oxide or the hydroxide of the metal with the appropriate acid.

Example :

zinc oxide	+	sulphuric acid (dil)	$\longrightarrow$	zinc sulphate	+	water
ZnO	+	H_2SO_4	$\longrightarrow$	$ZnSO_4$	+	H_2O
zinc hydroxide	+	sulphuric acid (dil)	$\longrightarrow$	zinc sulphate	+	water
$Zn(OH)_2$	+	H_2SO_4	$\longrightarrow$	$ZnSO_4$	+	$2H_2O$
lead (II) oxide	+	nitric acid (dil)	$\longrightarrow$	lead(II)nitrate	+	water
PbO	+	$2HNO_3$	$\longrightarrow$	$Pb(NO_3)_2$	+	H_2O
lead (II) hydroxide	+	nitric acid (dil)	$\longrightarrow$	lead(II)nitrate	+	water
$Pb(OH)_2$	+	$2HNO_3$	$\longrightarrow$	$Pb(NO_3)_2$	+	$2H_2O$

Method 4 Action of an acid on a carbonate

The carbonate of any metal can react with acids to give the corresponding salt of the metal, water and carbon dioxide.

zinc carbonate	+	sulphuric acid (dil)	$\longrightarrow$	zinc sulphate	+	water	+	carbon dioxide
$ZnCO_3$	+	H_2SO_4	$\longrightarrow$	$ZnSO_4$	+	H_2O	+	CO_2

calcium carbonate	+	hydrochloric acid (dil)	$\longrightarrow$	calcium chloride	+	water	+	carbon dioxide
$CaCO_3$	+	$2HCl$	$\longrightarrow$	$CaCl_2$	+	H_2O	+	CO_2
lead(II) carbonate	+	nitric acid (dil)	$\longrightarrow$	lead (II) nitrate	+	water	+	carbon dioxide
$PbCO_3$	+	$2HNO_3$	$\longrightarrow$	$Pb(NO_3)_2$	+	H_2O	+	CO_2

This method has its limitations. The reaction is incomplete if the carbonate is insoluble in water and, by its reaction with the acid, produces a salt which is also insoluble. In this case, the salt which is formed precipitates on the unchanged carbonate and stops the action.

For example, if dilute sulphuric acid is added to marble, $CaCO_3$, there is rapid effervescence for a few seconds, but the action quickly stops.

calcium carbonate (marble)	+	sulphuric acid (dil)	$\longrightarrow$	calcium sulphate	+	water	+	carbon dioxide
$CaCO_3$	+	H_2SO_4	$\longrightarrow$	$CaSO_4$	+	H_2O	+	CO_2
(insoluble)				(insoluble)				

The very slightly soluble calcium sulphate is precipitated on the marble making the calcium carbonate inside inaccessible to the acid and the reaction is stopped.

Method 5 Direct combination of two elements

Certain binary salts can be prepared by direct combination of the two constituent **elements.**

Example:

iron	+	**chlorine**	$\longrightarrow$	**iron (III) chloride**
2Fe	+	$3Cl_2$	$\longrightarrow$	$2FeCl_3$
iron	+	**sulphur**	$\longrightarrow$	**iron (II) sulphide**
Fe	+	**S**	$\longrightarrow$	**FeS**

Some common acids and bases

	where you find it and why you use it
Acids	
- Carbonic acid	- In fizzy drinks (carbonated drinks-releases bubbles of gas hissing sound)
- Citric acid	- In fruit juices
- Ethanoic acid (acetic acid)	- In vinegar
- Hydrochloric acid	- Used in "pickling" metals before they are coated
- Nitric acid	- Used for making fertilizers and explosives
- Phosphoric acid	- In rust inhibitor; used for making fertilizers
- Sulphuric acid	- In car batteries; used for making fertilizers
Bases	
- Ammonia	- In cleaning fluids as a degreasing agent; used in the manufacture of fertilizers
- Calcium hydroxide	- Spread on soil which is too acidic
- Calcium oxides	- Used in the manufacture of cement, mortar and concrete
- Magnesium hydroxide	- In 'antiacid' indigestion tablets and milk of magnesia
- Sodium hydroxide	- In powerful oven cleaners as a degreasing agent; used in the manufacture of soap

SUMMARY

In this chapter, the students would be able to study about the theories of acids, and bases, the principle of salt formation, properties and their prepartation methods. At the end of this chapter, questions and problems are given for exercise and practice.

Questions and Problems

1. Give the names and formulae of
   (a) an acid salt
   (b) a hydrated salt
   (c) a basic salt
2. Describe how you would prepare zinc sulphate from zinc carbonate.
3. Describe how you would obtain copper (II) sulphate crystals from copper (II) oxide.
4. 416 g anhydrous barium chloride were obtained when 488 g of the hydrated salt were heated. Calculate n in the formula $BaCl_2.nH_2O$
5. Explain what would happen when each of the following is heated.
   (a) Lead (II) carbonate
   (b) Lead (II) nitrate
6. Write TRUE or FALSE for each of the following statements.
   (a) An acid is a compound which, when dissolved in water, produces hydrogen ions, H^+ as the only positive ion.
   (b) Most acids turn red litmus blue.
   (c) A base will react with an acid to form salt and water.
   (d) Acids and bases cannot react with each other.
   (e) An acid reacts with a metal to produce soluble salt and hydrogen.
   (f) Most acids turn blue litmus red.
7. Fill in the blanks with a suitable word or words.
   (a) An acid when dissolved in water produces .... ions.
   (b) Most dilute acids have .... taste.
   (c) A base will usually react with an acid forming a .... and water only.
   (d) A normal salt is made up of a .... radical united with an acid radical.
   (e) Soluble salts are usually prepared by methods which involve ....
   (f) A base is a .... oxide or hydroxide which will react with an acid to give a salt and water.
   (g) A .... is a compound, produced from the reaction of a base or metal with an acid .
8. Select the correct word or words given in the bracket.
   (a) (Decomposition, Crystallization, Neutralization) is the reaction between an acid and a base to produce a salt and water.
   (b) Almost all acids react with carbonates to liberate (oxgen, hydrogen, carbon dioxide).
   (c) An alkali is a (basic, acidic, neutral) hydroxide which is soluble in water.

(d) An acid salt contains (exchangeable, replaceable, non-replaceable) hydrogen.
(e) The hydroxides of Al, Zn, Pb are (amphoteric, acidic, basic).
(f) (Calcium hydroxide, Copper (II) hydroxide, Zinc hydroxide) is an alkali.
(g) Which of the following statement(s) is(are) true for acid?
  (i) They turn red litmus blue.
  (ii) They react with carbonate to give a salt, water and carbon dioxide.
  (iii) They react with bases, forming salt and water.

9. Match each of the items given in **list A** with the appropriate item given in **list B.**

List A	**List B**
(a) An acid salt	(i) neutralization
(b) A base is usually a metallic oxide or hydroxide reacts with acid	(ii) contain neither replaceable hydrogen nor a hydroxyl group
(c) The reaction between an acid and a base	(iii) contains a replaceable hydrogen
(d) A normal salt	(iv) to produce a salt and water only
(e) An acid dissolving in water producing charged particles called ions.	(v) This process is known as dissociation

10. Define the following terms.
(a) An acid (c) A normal salt
(b) A base (d) An acid salt

11. Explain the following terms.
(a) Neutralization
(b) Crystallization
(c) Double decomposition

12. Name three common acids and bases.

****

CHAPTER 15

CARBON AND ITS COMPOUNDS

15.1 Occurrence

Carbon is found in the free state as diamond and graphite. Carbon in combination with other elements is found in petroleum, coal, natural gas and limestone. Also, all living things such as animals and plants contain carbon. The atmosphere contains 0.03% by volume of carbon dioxide.

15.2 Allotropy and Allotropes of Carbon

If an element, in the same physical state, can exist in more than one form, it is said to exhibit allotropy or polymorphism. The different forms in the same physical state possess different physical properties and may have different chemical properties. The differences between the forms may be due to :

(1) differences in crystal structure (diamond and graphite),

(2) differences in the number of atoms in the molecules as in the case of gases (oxygen and ozone), or

(3) differences in the molecular structure of the liquid forms (liquid sulphur at different temperatures).

Physical properties of crystalline allotropes of carbon

The physical properties of diamond and graphite which are the crystalline allotropes of carbon, are given in **Table 15.1.**

Table 15.1 Physical properties of diamond and graphite

Physical property	*Graphite*	*Diamond*
Density	2.3 g cm^{-3}	3.5 g cm^{-3}
Colour	grey-black	colourless
Opacity	opaque	transparent
Hardness	soft	very hard
Melting point	sublimes at about 3700°C	3550 °C

Structure of crystalline allotropes of carbon

● each dot represents a carbon atom

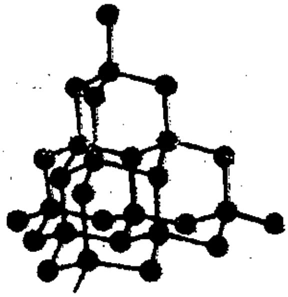

Fig. 15.1 The structure of diamond

In diamond, each carbon atom is surrounded by four other carbon atoms. Diamond has a gaint structure. Diamond contains millions of carbon atoms in a three dimensional network(Fig. 15.1). Therefore, diamond is very hard and has a very high melting point. (Diamond is the hardest among all naturally occuring substances.)

• each dot represent a carbon atom

1 nm 10^{-9} metre

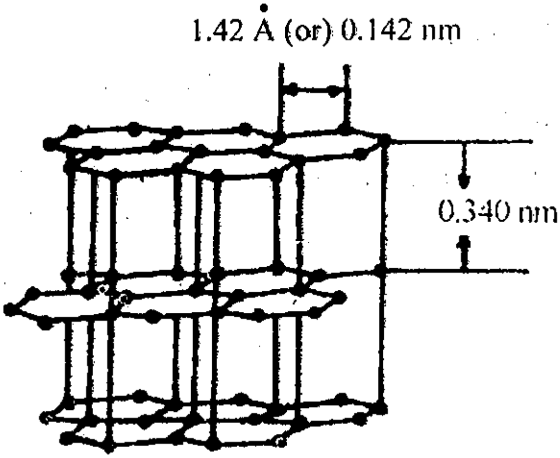

Fig. 15.2 The sturcture of graphite

In graphite, each carbon atom is surrounded by three other carbon atoms in the same plane and therefore layers of hexagons are obtained (Fig. 15.2). The distance between the layers is more than the distance between adjacent carbon atoms and so the layers are weakly bonded to each other. Therefore, graphite is soft.

Naturally occurring graphite is not enough for industrial use. Hence, to supplement industrial requirements, artificial graphite is prepared by heating powdered coke at high temperatures in an electric furnace.

To show that diamond and graphite are of the same element carbon

When carbon burns in oxygen, carbon dioxide is formed. Carbon dioxide dissolves in aqueous potassium hydroxide to form aqueous potassium carbonate.

carbon + oxygen $\longrightarrow$ carbon dioxide

$$C + O_2 \longrightarrow CO_2$$

potassium hydroxide + carbon dioxide $\longrightarrow$ potassium carbonate + water

$$2KOH + CO_2 \longrightarrow K_2CO_3 + H_2O$$

Since diamond and graphite are of the same element carbon, equal masses of diamond and graphite are found to produce equal masses of carbon dioxide.

Diamond and graphite may be burnt to form carbon dioxide under suitable experimental conditions. From such experimental determination, it has been definitely shown that the diamond and graphite are allotropic forms of the same element, carbon.

Other forms of carbon

Coal, coke, charcoal and carbon black are assumed to be amorphous forms of carbon. Now it is found that these forms of carbon contain randomly oriented small crystals of graphite.

Charcoal: Charcoal is made by heating wood in the absence of air. Charcoal is a black porous solid.

Coal: Coal is found in nature. In coal the element carbon is mixed with compounds of other elements. Coal is a black heavy solid.

Coke: Coke is formed by heating coal in the absence of air. Coke is a black heavy solid.

Carbon black (soot) : When kerosene is burnt in a limited amount of air, hydrogen from kerosene combines with oxygen from the air the carbon is left as carbon black. Carbon black is a black powder.

15.3 Uses of Carbon

Diamond

Cut diamonds are used for jewellery. Small pieces of diamonds are used as glass cutters and drill points.

Graphite

Graphite is used for making electrodes and crucibles. Graphite crucibles are used as containers for melting metals at high temperature. Graphite mixed with clay is used as pencil leads.

Charcoal

Charcoal has a porous structure and has many small holes. Because of this porosity, charcoal has a relatively larger surface area, and can remove colouring matter and gases by a process known as adsorption. This adsorptive property of charcoal can be further improved by treatment with dilute acids and other chemicals. Charcoal so treated is known as activated charcoal or activated carbon.

Activated charcoal is used in industry for bleaching (removal of colour) and deodourisation (removal of smell) of substances. Charcoal is also used as a fuel for cooking.

Coal

Coal is used to produce coke and coal tar.

Coke

Coke is used as a fuel in industry and as reducing agent in the extraction of metals, such as iron, lead, zinc etc.

Carbon Black

Carbon black is used form making printing ink, black shoe polish and as filler in vehicle tyres and other rubber products.

15.4 Chemical Properties of Carbon

1. When carbon burns in excess air or oxygen, carbon dioxide is formed.

$$\text{carbon} + \text{oxygen} \longrightarrow \text{carbon dioxide}$$

$$2C + O_2 \longrightarrow CO_2$$

Carbon monoxide is formed when carbon burns in a limited amount of air or oxygen.

$$\text{carbon} + \text{oxygen} \longrightarrow \text{carbon monoxide}$$

$$C + O_2 \longrightarrow 2CO$$

2. Red hot carbon combines with sulphur vapour to form carbon disulphide.

carbon + sulphur ⟶ carbon disulphide

$C + 2S \longrightarrow CS_2$

3. **Reducing properties of carbon**

When strongly heated, carbon can reduce the oxides of zinc and the metals below it in the activity series to their respective metals.

carbon + zinc oxide ⟶ zinc + carbon monoxide

$C + ZnO \longrightarrow Zn + CO$

carbon + lead (II) oxide ⟶ lead + carbon monoxide

$C + PbO \longrightarrow Pb + CO$

Carbon + coper (II) oxide ⟶ copper + carbon monoxide

$C + CuO \longrightarrow Cu + CO$

Red hot carbon reduces carbon dioxide to carbon monoxide.

carbon + carbon dioxide ⟶ carbon monoxide

$C + CO_2 \longrightarrow 2CO$

Red hot coke can reduce steam to hydrogen.

carbon + steam ⟶ carbon monoxide + hydrogen

$C + H_2O \longrightarrow CO + H_2$

4. Carbon is slowly oxidized to carbon dioxide by strong oxidizing acids such as concentrated sulphuric acid.

carbon + sulphuric acid $\xrightarrow{\Delta}$ carbon dioxide + sulphur dioxide + water

$C + 2H_2SO_4 \xrightarrow{\Delta} CO_2 + 2SO_2 + 2H_2O$

15.5 Compounds of Carbon

Oxides of carbon

There are two oxides of carbon, namely, carbon monoxide and carbon dioxide.

Carbon monoxide CO

Laboratory preparation of carbon monoxide from methanoic acid (formic acid)

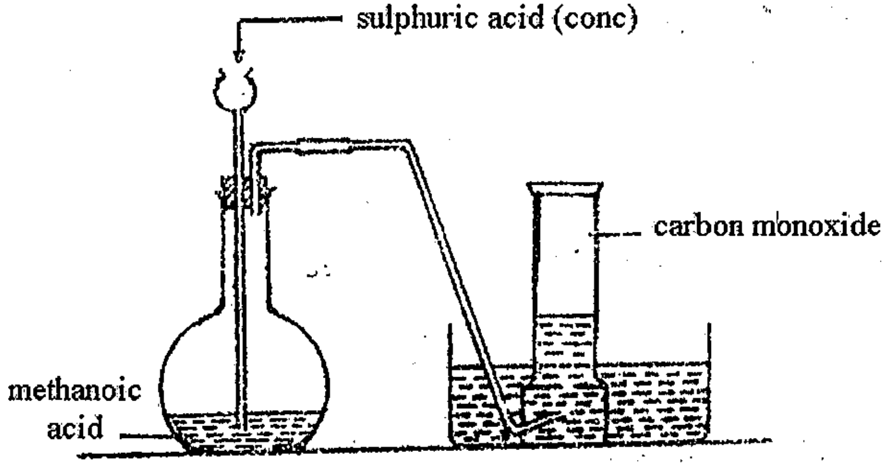

Fig. 15.3 Preparation of carbon monoxide from methanoic acid

$$\text{methanoic acid} \xrightarrow{H_2SO_4\text{(conc)}} \text{water} + \text{carbon monoxide}$$

$$HCOOH \xrightarrow{H_2SO_4\text{(conc)}} H_2O + CO$$

Some methanoic acid is placed in a flat-bottomed flask. and the apparatus set up as shown in **Fig. 15.3**. Concentrated sulphuric acid is slowly poured into the flask through a thistle funnel. Carbon monoxide is formed by the action of concentrated sulphuric acid on methanoic acid. Since carbon monoxide is only slightly soluble in water and has about the same relative vapour density as air, it is collected by the downward displacement of water. The gas jar is filled with carbon monoxide, when all the water in the gas jar is displaced.

Physical properties of carbon monoxide

Carbon monoxide is a colourless, odourless gas. It is slightly soluble in water, but this solution does not change the colour of litmus paper. Carbon monoxide is very poisonous. (Carbon monoxide can combine with haemoglobin of the blood and thus prevent the blood from carrying oxygen for use in the body.) Carbon monoxide has about the same relative vapour density as air.

Chemical properties of carbon monoxide

1. Carbon monoxide burns with a blue flame in air.

carbon monoxide + oxygen ⟶ carbon dioxide

$$2CO + O_2 \longrightarrow 2CO_2$$

2. **Reducing properties of carbon monoxide**

Carbon monoxide can reduce the oxides of zinc and the metals below it in the activity series to the respective metals.

Example:

When carbon monoxide is passed over heated iron (III) oxide, metallic iron is formed.

iron(III) oxide + carbon monoxide ⟶ iron + carbon dioxide

$$Fe_2O_3 + 3CO \longrightarrow 2Fe + 3CO_2$$

Carbon monoxide can also reduce heated copper (II) oxide to metallic copper.

copper(II) oxide + carbon monoxide ⟶ copper + carbon dioxide

$$CuO + CO \longrightarrow Cu + CO_2$$

Heated lead (II) oxide is reduced to metallic lead by carbon monoxide.

lead(II) oxide + carbon monoxide ⟶ lead + carbon dioxide

$$PbO + CO \longrightarrow Pb + CO_2$$

Carbon monoxide can reduce heated zinc oxide to metallic zinc.

zinc oxide + carbon monoxide ⟶ zinc + carbon dioxide

$$ZnO + CO \longrightarrow Zn + CO_2$$

3. Addition properties of carbon monoxide

When a mixture of carbon monoxide and sulphur vapour is passed through a heated tube, carbonyl sulphide is formed.

sulphur + carbon monoxide ⟶ carbonyl sulphide

$$S + CO \longrightarrow COS$$

Carbonyl chloride is formed when a mixture of carbon monoxide and chlorine is exposed to sunlight.

chloride + carbon monoxide ⟶ carbonyl chloride

$$Cl_2 + CO \longrightarrow COCl_2$$

Carbon monoxide combines with bromine vapour to form carbonyl bromide.

bromine + carbon monoxide ⟶ carbonyl bromide

$$Br_2 + CO \longrightarrow COBr_2$$

Carbon monoxide combines with gently heated nickel to form nickel carbonyl.

nickel + carbon monoxide ⟶ nickel carbonyl

$$Ni + 4CO \longrightarrow Ni(CO)_4$$

When sodium hydroxide is heated to 200°C and carbon monoxide is passed over it, sodium methanoate is obtained.

sodium hydroxide + carbon monoxide ⟶ sodium methanoate

$$NaOH + CO \longrightarrow HCOONa$$

A solution of copper (I) chloride in concentrated hydrochloric acid or ammonia solution can absorb carbon monoxide gas.

copper (I) chloride + water + carbon monoxide ⟶ copper (I) chloride carbon monoxide complex

$$CuCl + 2H_2O + CO \longrightarrow CuCl.\ CO.\ 2H_2O$$

Carbon dioxide CO_2

Laboratory preparation of carbon dioxide from calcium carbonate

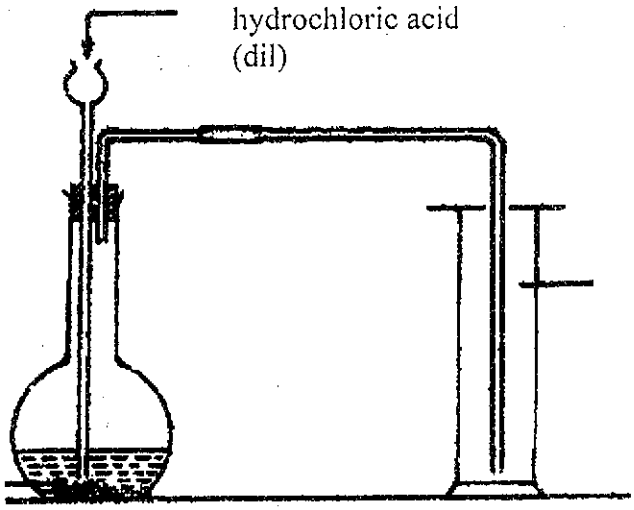

Fig.15.4 Preparation of carbon dioxide from calcium carbonate

calcium carbonate	+	hydrochloric acid (dil)	$\longrightarrow$	calcium chloride	+	water	+	carbon dioxide
$CaCO_3$	+	$2HCl$	$\longrightarrow$	$CaCl_2$	+	H_2O	+	CO_2

Some marble chips (calcium carbonate)are placed in a flat-bottomed flask and the apparatus set up as shown in Fig. 15.4. Dilute hydrochloric acid is poured into the flask thorugh the thistle funnel. Carbon dioxide, being heavier than air, is collected by the upward displacement of air. Since carbon dioxide is a colourless, odourless gas, testing is required to establish whether the gas jar is full of carbon dioxide or not. If a burning candle, brought near the mouth of the gas jar, is extinguished, the gas jar is full of carbon dioxide.

Physical properties of carbon dioxide

Carbon dioxide is a colourless, odourless, tasteless gas. It is sparingly soluble in water and heavier than air. Carbon dioxide solidifies directly when cooled at atmoshperic pressure to give dry ice, i.e., solid carbon dioxide sublimes at - 78°C.

Chemical properties of carbon dioxide

1. Carbon dioxdie does not burn and does not support combustion (Fig. 15.5). However, the temperature of burning magnesium ribbon is high enough to decompose the carbon dioxide and liberate oxygen. The magnesium ribbon continues to burn in the presence of this liberated oxygen.

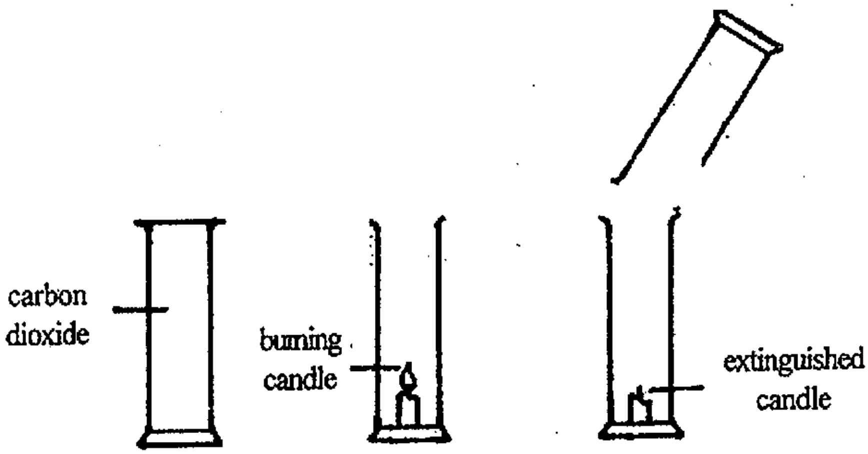

Fig. 15.5 Extinguishing a candle with carbon dioxide

magnesium	+	carbon dioxide	$\longrightarrow$	magnesium oxide	+	carbon
$2Mg$	+	CO_2	$\longrightarrow$	$2MgO$	+	C

2. Acidic properties of carbon dioxide

Carbon dioxide is sparingly soluble in water forming a weakly acidic solution of carbonic acid which turns blue litmus pink.

carbon dioxide	+	water	$\longrightarrow$	carbonic acid
CO_2	+	H_2O	$\longrightarrow$	H_2CO_3

Carbon dioxide can react with basic oxides.

calcium oxide	+	carbon dioxide	$\longrightarrow$	calcium carbonate
CaO	+	CO_2	$\longrightarrow$	$CaCO_3$
sodium oxide	+	carbon dioxide	$\longrightarrow$	sodium carbonate
Na_2O	+	CO_2	$\longrightarrow$	Na_2CO_3

Carbon dioxide turns lime water milky due to the formation of insoluble calcium carbonate.

calcium hydroxide + carbon dioxide ⟶ calcium carbonate + water

$$Ca(OH)_2 + CO_2 \longrightarrow CaCO_3 + H_2O$$

Calcium carbonate redissloves on passing excess carbon dioxide to give a solution of calcium hydrogencarbonate.

calcium carbonate + water + carbon dioxide ⟶ calcium hydrogencarbonate

$$CaCO_3 + H_2O + CO_2 \longrightarrow Ca(HCO_3)_2$$

Uses of carbon dioxide

Since carbon dioxide is about one and a half times as dense as air and neither supports combustion nor is inflammable, it is used in fire extinguishers. The heavy carbon dioxide gas forms a blanket over the burning material, cuts off the air supply and puts out the flames.

Carbon dioxide is used in the preparation of mineral waters. Soda-water is a solution of carbon dioxide in water under pressure. Sweetened and flavoured soda-water is sold as lemonade.

Carbonates

Carbonic acid is a week dibasic acid and so it can form two kinds of salts depending on whether one or both of the hydrogens are replaced by a metal.

$$\begin{array}{ccc} & H_2CO_3 & \\ \swarrow & & \searrow \\ MHCO_3 & & M_2CO_3 \\ \text{hydrogencarbonate} & & \text{carbonate} \end{array}$$

M= a mono valent metal

Preparation of hydrogencarbonates

Hydrogencarbonates are prepared by passing carbon dioxide through solutions or suspensions of carbonates in water.

Example :

When carbon dioxide is passed through the solution of sodium carbonate in water, sodium hydrogencarbonate solution is formed.

sodium carbonate + water + carbon dioxide ⟶ sodium hydrogencarbonate

$$Na_2CO_3 + H_2O + CO_2 \longrightarrow 2NaHCO_3$$

When carbon dioxide is passed through the suspension of calcium carbonate in water, aqueous calcium hydrogencarbonate is formed.

calcium carbonate + water + carbon dioxide ⟶ calcium hydrogencarbonate

$$CaCO_3 + H_2O + CO_2 \longrightarrow Ca(HCO_3)_2$$

Properties of hydrogencarbonates

1. All hydrogencarbonates are soluble in water. Only the hydrogencarbonates of potassium, sodium and ammonium can exist in the solid state. Solutions of other hydrogencarbonate decompose on heating for **crystallization**.

calcium hydrogencarbonate ⟶ calcium carbonate + water + carbon dioxide

$$Ca(HCO_3)_2 \longrightarrow CaCO_3 + H_2O + CO_2$$

2. **All hydrogencarbonates decompose on heating.**

sodium hydrogencarbonate ⟶ sodium carbonate + water + carbon dioxide

$$2NaHCO_3 \longrightarrow Na_2CO_3 + H_2O + CO_2$$

magnesium hydrogencarbonate ⟶ magnesium carbonate + water + carbon dioxide

$$Mg(HCO_3)_2 \longrightarrow MgCO_3 + H_2O + CO_2$$

3. **All hydrogencarbonates react with dilute acids to give carbon dioxide.**

sodium hydrogencarbonate + hydrochloric acid ⟶ sodium chloride + water + carbon dioxide

$$NaHCO_3 + HCl \longrightarrow NaCl + H_2O + CO_2$$

Preparation of carbonates

Some of the carbonates such as sodium carbonate, calcium carbonate and magnesium carbonate occur in nature. Sodium carbonate is used for the preparation of other carbonates.

When an aqueous solution of sodium carbonate is mixed with an aqueous solution of metal salt, either normal carbonate or basic carbonate is precipitated.

When aqueous sodium carbonate is mixed with aqueous calcium chloride, calcium carbonate is precipitated.

sodium carbonate	+	calcium chloride	$\longrightarrow$	calcium carbonate	+	sodium cloride
Na_2CO_3	+	$CaCl_2$	$\longrightarrow$	$CaCO_3$	+	$2NaCl$

When aqueous sodium carbonate is mixed with aqueous lead (II) nitrate, basic lead (II) carbonates precipitated.

The various stages of the reaction may be represented as follows:

lead (II) nitrate	+	sodium carbonate	$\longrightarrow$	lead (II) carbonate	+	sodium nitrate
$Pb(NO_3)_2$	+	Na_2CO_3	$\longrightarrow$	$PbCO_3$	+	$2NaNO_3$
sodium carbonate	+	water	$\longrightarrow$	sodium hydroxide	+	carbonic acid
Na_2CO_3	+	$2H_2O$	$\longrightarrow$	$2NaOH$	+	H_2CO_3
Lead (II) nitrate	+	sodium hydroxide	$\longrightarrow$	lead (II) hydroxide	+	sodium nitrate
$Pb(NO_3)_2$	+	$2NaOH$	$\longrightarrow$	$Pb(OH)_2$	+	$2NaNO_3$
lead (II) carbonate	+	lead (II) hydroxide	$\longrightarrow$	dibasic lead (II) carbonate		
$2PbCO_3$	+	$Pb(OH)_2$	$\longrightarrow$	$2PbCO_3 . Pb(OH)_2$		

There are no basic carbonates of potassium, sodium and calcium. Calcium carbonate is therefore prepared by the reaction between and aqueous solution of soluble calcium salts and aqueous sodium carbonate. The carbonates of other heavy metals may be prepared by using sodium hydrogencarbonate instead of sodium carbonate.

lead(II) nitrate + sodium hydrogen carbonate $\longrightarrow$ lead(II) carbonate + sodium nitrate + carbon dioxide + water

$$Pb(NO_3)_2 + 2NaHCO_3 \longrightarrow PbCO_3 + 2NaNO_3 + CO_2 + H_2O$$

Carbonates are obtained when carbon dioxide is passed through aqueous solutions of hydroxides.

calcium hydroxide + carbon dioxide $\longrightarrow$ calcium carbonate + water

$$Ca(OH)_2 + CO_2 \longrightarrow CaCO_3 \downarrow + H_2O$$

Properties of carbonates

1. Carbonates of potassium, sodium and ammonium are colourless and soluble in water. Carbonates and basic carbonates of other metals are insoluble in water.

2. **Action of heat on carbonates**

Carbonates of sodium and potassium do not decompose on heating. Ammonium carbonate decomposes readily on heating to ammonia, carbon dioxide and water.

ammonium carbonate $\longrightarrow$ ammonia + water + carbon dioxide

$$(NH_4)_2CO_3 \longrightarrow 2NH_3 + H_2O + CO_2$$

Insoluble carbonates decompose into oxide and carbon dioxide on heating.

calcium carbonate $\longrightarrow$ calcium oxide + carbon dioxide

$$CaCO_3 \longrightarrow CaO + CO_2$$

magnesium carbonate $\longrightarrow$ magnesium oxide + carbon dioxide

$$MgCO_3 \longrightarrow MgO + CO_2$$

Action of acids on carbonates

All carbonates react with dilute acids to give salts, water and carbon dioxide.

calcium carbonate	+	hydrochloric acid (dil)	$\longrightarrow$	calcium chloride	+	water	+	carbon dioxide
$CaCO_3$	+	$2HCl$	$\longrightarrow$	$CaCl_2$	+	H_2O	+	CO_2

sodium carbonate	+	sulphuric acid (dil)	$\longrightarrow$	sodium sulphate	+	water	+	carbon dioxide
Na_2CO_3	+	H_2SO_4	$\longrightarrow$	Na_2SO_4	+	H_2O	+	CO_2

basic copper(II) carbonate	+	nitric acid (dil)	$\longrightarrow$	copper(II) nitrate	+	water	+	carbon dioxide
$Cu(OH)_2.\ 2CuCO_3$	+	$6HNO_3$	$\longrightarrow$	$3Cu(NO_3)_2$	+	$4H_2O$	+	$2CO_2$

Aqueous solution of potassium carbonate, sodium carbonate or ammonium carbonate, when mixed with metallic salt solution give either carbonate or basic carbonate.

Test for carbonates

Some of the sample is put in a test tube.

Dilute sulphuric acid is added.If a carbonate is present, there will be effervesence.The gas evolved is passed into another test tube containing lime water.The lime water turns milky, if the gas is carbon dioxide and the sample is carbonate.

sodium carbonate	+	sulphuric acid (dil)	$\longrightarrow$	sodium sulphate	+	Water	+	carbon dioxide
Na_2CO_3	+	H_2SO_4	$\longrightarrow$	Na_2SO_4	+	H_2O	+	CO_2

carbon dioxide	+	calcium hydroxide	$\longrightarrow$	calcium carbonate	+	Water
CO_2	+	$Ca(OH)_2$	$\longrightarrow$	$CaCO_3$	+	H_2O

The greenhouse effect

The earth's climate is effected by the levels of carbon dioxide and water vapour in the atmosphere. These gases are responsible for the **greenhouse effect.**

The earth is heated by the sun. Heat energy from the earth is radiated back into space. The balance keeps the temperature of the anything happens to upset the balance, the result will be a change in the world's climate.

A rise in temperature could melt the ice caps at the North and South poles. A fall could decrease the world's food production.

When the Earth sends heat energy back into space, it changes the wavelength to that of infrared radiation. Unlike sunlight, infrared radiation cannot travel freely through the air surrounding the Earth. Water vapour and carbon dioxide absorb infrared radiation. They act as blankets round the earth, hindering the escape of heat into space. Without these blankets the Earth's surface would be (-40°C), instead of a life-supporting (15°C). If additional water vapour and carbon dioxide were added to the atmosphere, it would be like adding another blanket. The Earth would heat up. This blanketing by water vapour and carbon dioxide is called **the greenhouse effect.**

SUMMARY

In this chapter ,carbon and its compounds are discussed in detail. Carbon can exist in many allotropic forms. Two crystalline forms of carbon are diamond and graphite. Amorphous forms of carbon are coal, coke , charcoal, and carbon black. Compounds of carbon such as carbon monoxide, and carbon dioxide, are elaborated with preparation methods, physical and chemical properties and uses.In addition, preparation , properties and tests for hydrogencarbonates and carbonates are also presented. The greenhouse effect caused by carbon dioxide and water vapour is summarized as follows.

The greenhouse effect

A layer of carbon dioxide and water vapour surrounds the earth. This layer reduces the amount of heat which the Earth radiates into space. By burning hydrocarbon fuels, we pour carbon dioxide and water vapour into the atmosphere. The layer of carbon dioxide and water vapour is growing thicker. If it becomes too thick, the earth will warm up, and the ice caps could begin to melt. This blanketing by water vapour and carbon dioxide is called **the greenhouse effect**. We also put many gases into the atmosphere. This has the effect of scattering some of the sun's radiation and cooling the earth.

Questions and Problems

1. How can you show that diamond and graphite are different forms of the same element carbon?

2. Even though diamond and graphite are the allotropes of carbon, explain why they have different physical properties.

3. Give five examples to show that carbon has reducing properties.

4. Carbon monoxide and carbon dioxide are colourless, odourless gases.

   How will you identify these gases in separate gas jars?

   (a) Do you think the carbon dioxide obtained by the method described on page 201 is pure?

   (b) If it is not pure, what is the likely impurity?

   (c) How would you remove this impurity and obtain pure dry carbon dioxide?

   (d) Draw the diagram of the apparatus you would use and describe the experiment.

5. Write equations in words and symbols for the following reactions:

   (a) $H_2SO_4(l) + C(s) \longrightarrow ?$

   (b) $Fe_2O_3(s) + CO(g) \longrightarrow ?$

   (c) $KOH(aq) + CO_2(g) \longrightarrow ?$

6. How would you differentiate aqueous sodium carbonate from aqueous sodium hydrogencarbonate?

7. Write TRUE or FALSE for each of the following statements.

   (a) Carbon is found in free state as diamond and graphite.

   (b) Carbon in combination with other element is found in petroleum coal, natural gas and limestone.

   (c) Carbon cannot exhibit allotropy or polymorphism.

   (d) Coal, coke, charcoal and carbon black are the amorphous forms of carbon.

(e) Diamond and graphite do not contain the same element carbon.

8. Fill in the blanks with a suitable word or words.

(a) Diamond and graphite are the .... allotropes of carbon.

(b) Coke, coal, charcoal and carbon black are the .... forms of carbon.

(c) Diamond and graphite are of the same element ....

(d) The two oxides of carbon are carbon monoxide and ....

(e) Carbon monoxide can also reduce heated copper (II) oxide to metallic.................

(f) Carbon dioxide is .... than air.

(g) Carbon dioxide solidifies when cooled to give ....

(h) Dry ice sublimes at ......^{0}C.

(i) Carbon dioxide can be used as a fire ....

9. Select the correct word or words given in the bracket.

(a) Diamond and graphite belong to (amorphous, crystalline, different) forms of carbon

(b) Carbon monoxide can (oxidise, reduce, promote) heated zinc oxide to metallic zinc.

(c) Coal, coke, charcoal, carbon black are (crystalline, amorphous, same) forms of carbon.

(d) All carbonates react with dilute acids to give salts, water and (carbon **dioxide, oxygen, sulphur dioxide).**

(e) When carbon monoxide is passed over heated iron (III) oxide, metallic (lead, copper, iron) is formed.

(f) Which of these is not an allotrope of carbon?

(i) coal (ii) graphite (iii) **charcoal**

(g) **Which of the following is not a natural form of calcium carbonate?**

(i) chalk (ii) lime (iii) limestone (iv) marble

(h) Which of the following properties does not apply to carbon dioxide?

(i) It is slightly soluble in water. (iii) It is a reducing agent.

(ii) It is denser than air. (iv) It is flammable.

(i) When carbon dioxide is bubbled into lime water , it turns milky. The white precipitate is due to the formation of insoluble ,

(i) calcium oxide (iii) calcium hydrogencarbonate

(ii) calcium hydroxide (iv) calcium carbonate

10. Match each of the items given in **list A** with the appropriate item given in **list B**.

	List A		**List B**
(a)	Diamond is the crystalline form of carbon	(i)	to form potassium carbonate
(b)	Diamond and graphite are the crystalline forms of carbon	(ii)	It has a giant structure
(c)	Carbon dioxide dissolves in aqueous potassium hydroxide	(iii)	equal masses of diamond and graphite are found to produce equal masses of CO_2
(d)	Carbon can exit in amorphous and crystalline forms	(iv)	They contain the same element carbon
(e)	Diamond and graphite are of the same element carbon	(v)	It exhibits allotropy or polymorphism

11. Explain the greenhouse effect.

12. Answer the following questions.

(a) Name two allotropes of carbon..

(b) Explain the following observations with suitable equations.

(i) When carbon dioxide is bubbled through lime water, a milky white suspension forms.

(ii) But if you continue to bubble through more carbon dioxide gas,the suspension disappears and form a clear liquid.

(c) Give equations for addition properties of carbon monoxide.

(d) Describe the main steps involved in the test for carbonate.

(e) What is polymorphism? Give example in crystal, in gases, in molecular structure.

****

PERIODIC TABLE OF THE ELEMENTS

Transition Elements

Group	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
Period	IA	IIA	IIIA	IVA	VA	VIA	VIIA	VIII	VIII	VIII	IB	IIB	IIIB	IVB	VB	VIB	VIIB	0
1	1 H 1.008																	2 He 4.003
2	3 Li 6.94	4 Be 9.01											5 B 10.82	6 C 12.011	7 N 14.008	8 O 16	9 F 19	10 Ne 20.183
3	11 Na 22.991	12 Mg 24.32											13 Al 26.98	14 Si 28.09	15 P 30.975	16 S 32.06	17 Cl 35.457	18 Ar 39.944
4	19 K 39.1	20 Ca 40.08	21 Sc 44.96	22 Ti 47.9	23 V 50.94	24 Cr 52.01	25 Mn 54.94	26 Fe 55.85	27 Co 58.94	28 Ni 58.71	29 Cu 63.54	30 Zn 65.38	31 Ga 69.72	32 Ge 72.59	33 As 74.91	34 Se 78.96	35 Br 79.916	36 Kr 83.8
5	37 Rb 85.47	38 Sr 87.63	39 Y 88.91	40 Zr 91.22	41 Nb 92.91	42 Mo 95.95	43 Tc 97	44 Ru 101.07	45 Rh 102.91	46 Pd 106.4	47 Ag 107.88	48 Cd 112.4	49 In 114.82	50 Sn 118.7	51 Sb 121.76	52 Te 127.6	53 I 126.91	54 Xe 131.3
6	55 Cs 132.91	56 Ba 137.36	57* La 138.91	72 Hf 178.49	73 Ta 180.95	74 W 183.86	75 Re 186.21	76 Os 190.2	77 Ir 192.22	78 Pt 195.09	79 Au 197	80 Hg 200.61	81 Tl 204.37	82 Pb 207.21	83 Bi 209	84 Po 209	85 At 210	86 Rn 222
7	87 Fr 223	88 Ra 226.05	89** Ac 227	104 Rf 261	105 Db 262	106 Sg 266	107 Bh [illegible]	108 Hs [illegible]	109 Mt 268	110 Ds 271	111 Uuu 272	112 Uub 277	113 Uut 284	114 Uuq 285	115 Uup 288	116 Uuh 292	117 Uus	118 Uuo

*Lanthanum Series	58 Ce 140.12	59 Pr 140.91	60 Nd 144.24	61 Pm 145	62 Sm 150.4	63 Eu 151.96	64 Gd 157.25	65 Tb 158.93	66 Dy 162.5	67 Ho 164.93	68 Er 167.26	69 Tm 168.93	70 Yb 173.04	71 Lu 174.97
**Actinium Series	90 Th 232.04	91 Pa 231.04	92 U 238.07	93 Np 237.05	94 Pu 244	95 Am 243	96 Cm 247	97 Bk 247	98 Cf 251	99 Es 254	100 Fm 257	101 Md 258	102 No 255	103 Lw 260

Alkali metals | Alkaline earth metals | Transition metals | Other metals | Nonmetals

Halogens | Noble gases | Lanthanum series | Actinium series | Darkline (Metalloids)

Table of relative atomic numbers and atomic masses of the elements

Element	Atomic Number	Symbol	Atomic mass	Element	Atomic Number	Symbol	Atomic mass
Actinium	89	Ac	227	Mercury	80	Hg	200.61
Aluminium	13	Al	26.98	Molybdenum	42	Mo	95.95
Americium	95	Am	243	Neodymium	60	Nd	144.24
Antimony	51	Sb	121.76	Neon	10	Ne	20.183
Argon	18	Ar	39.944	Neptunium	93	Np	237.05
Arsenic	33	As	74.91	Nickel	28	Ni	58.71
Astatine	85	At	210	Niobium	41	Nb	92.91
Barium	56	Ba	137.36	Nitrogen	7	N	14.008
Berkelium	97	Bk	247	Nobelium	102	No	255
Beryllium	4	Be	9.01	Osmium	76	Os	190.2
Bismuth	83	Bi	209.00	Oxygen	8	O	16
Boron	5	B	10.82	palladium	46	Pd	106.4
Bromine	35	Br	79.916	Phosphorus	15	P	30.975
Cadmium	48	Cd	112.40	Platinum	78	Pt	195.09
Calcium	20	Ca	40.08	Plutonium	94	Pu	244
Californium	98	Cf	251	Polonium	84	Po	209
Carbon	6	C	12.011	Potassium	19	K	39.100
Cerium	58	Ce	140.12	Praseodymium	59	Pr	140.91
Caesium	55	Cs	132.91	Promethium	61	Pm	145
Chlorine	17	Cl	35.457	Protactinium	91	Pa	231.04
Chromium	24	Cr	52.01	Radium	88	Ra	226.05
Cobalt	27	Co	58.94	Radon	86	Rn	222
Copper	29	Cu	63.54	Rhenium	75	Re	186.21
Curium	96	Cm	247	Rhodium	45	Rh	102.91
Dysprosium	66	Dy	162.5	Rubidium	37	Rb	85.47
Einsteinium	99	Es	254	Ruthenium	44	Ru	101.07
Erbium	68	Er	167.26	Samarium	62	Sm	150.4
Europium	63	Eu	151.96	Scandium	21	Sc	44.96
Fermium	100	Fm	257	Selenium	34	Se	78.96
Fluorine	9	F	19.00	Silicon	14	Si	28.09
Francium	87	Fr	223	Silver	47	Ag	107.880
Gadolinium	64	Gd	157.25	Sodium	11	Na	22.991
Gallium	31	Ga	69.72	Strontium	38	Sr	87.63

Element	Atomic Number	Symbol	Atomic mass	Element	Atomic Number	Symbol	Atomic mass.
Germanium	32	Ge	72.59	Sulphur	16	S	32.066
Gold	79	Au	197.0	Tantalum	73	Ta	180.95
Hafnium	72	Hf	178.49	Technetium	43	Tc	97
Helium	2	He	4.003	Tellurium	52	Te	127.60
Holmium	67	Ho	164.93	Terbium	65	Tb	158.93
Hydrogen	1	H	1.008	Thallium	81	Tl	204.37
Indium	49	In	114.82	Thorium	90	Th	232.04
Iodine	53	I	126.91	Thulium	69	Tm	168.93
Iridium	77	Ir	192.22	Tin	50	Sn	118.70
Iron	26	Fe	55.85	Titanium	22	Ti	47.90
Krypton	36	Kr	83.80	Tungsten (Wolfram)	74	W	183.86
Lanthanum	57	La	138.91	Uranium	92	U	238.07
Lawrencium	103	Lr	260	Vanadium	23	V	50.94
Lead	82	Pb	207.21	Xenon	54	Xe	131.30
Lithium	3	Li	6.940	Ytterbium	70	Yb	173.04
Lutetium	71	Lu	174.97	Yttrium	39	Y	88.91
Magnesium	12	Mg	24.32	Zinc	30	Zn	65.38
Manganese	25	Mn	54.94	Zirconium	40	Zr	91.22
Mendelevium	101	Md	258				

Discovery of New Elements

Sr	Atomic No.	Symbol	Name	Atomic Wt	Electronic Structure	Metal	Discovered Year
1	104	Rf	Rutherfordium	261	2.8.18.32.32.10.2	Transition	1969
2	105	Db	Dubnium	262	2.8.18.32.32.11.2	Transition	1970
3	106	Sg	Seaborgium	266	2.8.18.32.32.12.2	Transition	1974
4	107	Bh	Bohnium	264	2.8.18.32.32.13.2	Transition	1976
5	108	Hs	Hassium	263	2.8.18.32.32.14.2	Transition	1984
6	109	Mt	Meitnerium	268	2.8.18.32.32.15.2	Transition	1982
7	110	Ds	Darnstadtium	271	2.8.18.32.32.17.2	Transition	1994
8	111	Uuu	Unununium	272	2.8.18.32.32.18.1	Transition	1994
9	112	Uub	Ununbium	277	2.8.18.32.32.18.2	Transition	1996
10	113	Uut	Ununtrium	284	2.8.18.32.32.18.3	Other elements	2003
11	114	Uuq	Ununquadium	285	2.8.18.32.32.18.4	Other elements	1998
12	115	Uup	Ununpentium	288	2.8.18.32.32.18.5	Other elements	2003
13	116	Uuh	Ununhexium	292	2.8.18.32.32.18.6	Other elements	2000
14	117	Uus	Ununseptium	-	-	Other elements	-
15	118	Uuo	Ununoctium	-	-	Other elements	-

Vapour Pressure of Water

Temperature °*C*	*Pressure* *mm Hg*	*Temperature* °*C*	*Pressure* *mm Hg*
0	4.6	23	21.2
1	4.9	24	22.4
2	5.3	25	23.8
3	5.7	26	25.2
4	6.1	27	26.7
5	6.5	28	28.3
6	7.0	29	30.0
7	7.5	30	31.8
8	8.0	35	42.2
9	8.6	40	55.3
10	9.2	45	71.9
11	9.8	50	92.5
12	10.5	55	118.0
13	11.2	60	149.4
14	12.0	65	187.5
15	12.8	70	233.7
16	13.6	75	289.1
17	14.5	80	355.1
18	15.5	85	433.6
19	16.5	90	525.8
20	17.5	95	633.9
21	18.7	100	760.0
22	19.8	105	906.1

The Activity Series **Appendix 1**

Element	Electro-positivity	Action of water	Action of acids	Action of air	Action of H_2 on heated oxide	Action of C on heated oxide	Nature of hydroxide	Action of heat on nitrate	Occurs naturally as
K		Violent at room temperature	Explosive. Gives off H_2	Tarnishes very rapidly KOH K_2CO_3					Chloride
Na		Vigorous at room temperature	Violent. Gives off H_2	Tarnishes very rapidly NaOH Na_2CO_3			Strong bases stable to heat	Give the nitrite and oxygen	Chloride
Ca		Steady reaction with cold water	Vigorous reaction giving H_2	Tarnishes very rapidly CaO $Ca(OH)_2$ $CaCO_3$	No reaction	Give the carbide at high temperature	Strong base, gives oxide on heating		Chloride and carbonate
Mg		Reacts with steam					Weak base, gives oxide on heating		Chloride and carbonate
Al	Steadily decreases	Red hot metal reacts with steam		Become coated with a protective layer of oxide					Oxide
Zn			Gives off H_2				Amphoteric. Gives oxide on heating		Oxide and sulphide
Fe		Reversible		Rusts slowly	Reversible reduction		Weak base, gives oxide on heating	Give oxide, nitrogen dioxide and oxygen	Oxide and sulphide
Pb					Rapid reduction	Reduction to the metal	Amphoteric. Gives oxide on heating		Sulphide
H				Oxidize readily on treating			Neutral		Oxide
Cu		No reduction	Do not liberate H_2		Rapid reduction		Weak base, gives oxide on heating		Sulphide
Ag			Attacked only by oxidizing acids	Attacked only by traces of sulphur compounds	Oxide decomposes on heating without the need for H_2	Decomposes by heat alone	Unstable at room temperature	Give the metal, nitrogen dioxide and oxygen	Sulphide and as free metal