(a) Symbols

A symbol is a shorthand notation of an element. The first letter in the name of an element is usually chosen as the symbol of the element. 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. A symbol represents an atom of the element and the mass of an atom.

Symbols of some elements

The symbols of some elements are taken from English names. The symbols of some other elements, which have been known since earlier times, are taken from Latin names (Tables 4.1 and 4.2).

Table 4.1 Symbols of Some Metallic Elements

English name	Latin name	Symbol
Sodium	Natrium	Na
Potassium	Kalium	K
Iron	Ferrum	Fe
Tin	Stannum	Sn
Lead	Plumbum	Pb
Copper	Cuprum	Cu
Mercury	Hydrargyrum	Hg
Silver	Argentum	Ag
Gold	Aurum	Au
Antimony	Stibium	Sb

English name	Symbol
Calcium	Ca
Barium	Ba
Magnesium	Mg
Aluminium	Al
Manganese	Mn
Zinc	Zn
Chromium	Cr
Platinum	Pt

Table 4.2 Symbols of Some Non- Metallic Elements

Name	Symbol
Carbon	C
Sulphur	S
Phosphorus	P
Oxygen	O
Hydrogen	H

Name	Symbol
Nitrogen	N
Chlorine	Cl
Bromine	Br
Iodine	I

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(b) Formulae

(i) Molecular formulae of elements

A molecular formula is a shorthand notation representing a molecule or a compound of a substance. It shows the number of atoms contained in a molecule or a compound.

A molecular formula also represents a molecule of the corresponding elements. **For example**, Cl_2 means a molecule of chlorine. $2O_2$ means two molecules of oxygen.

(ii) 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.

A molecule of carbon monoxide contains an atom of carbon and an atom of oxygen. The molecular formula of carbon monoxide is CO. A molecule of ammonia consists of an atom of nitrogen and three atoms of hydrogen. The molecular formula of ammonia is NH_3.

(iii) 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.

For 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 NaCl.

Similarly, the formula of magnesium oxide is MgO.

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.

(iv) Empirical formula

The first step in identifying a chemical formula is to find out its empirical formula. The empirical formula of a compound is the simplest formula of the compound. It shows

- the types of elements present in the compound;
- the simplest ratio of the different types of the atoms in the compound.

Example 1: When 1.55 g of phosphorus is completely combusted, 3.55 g of an oxide of phosphorus is produced. Deduce the empirical formula of this oxide of phosphorus. (O = 16, P = 31)

	P	O
Step 1 the mass of each element	1.55	3.55 – 1.55 = 2.00
Step 2 divide by relative atomic masses	$\frac{1.55}{31} = 0.05$	$\frac{2.00}{16} = 0.125$
Step 3 divide by the lowest number	$\frac{0.05}{0.05} = 1$	$\frac{0.125}{0.05} = 2.5$
	$1 \times 2 = 2$	$2.5 \times 2 = 5$
Step 4 empirical formula (must be in integers)	P_2O_5	

An empirical formula can also be deduced from data that give the percentage composition by mass of the elements in a compound.

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Example 2: A compound of carbon and hydrogen contains 85.7 % of carbon and 14.3 % of hydrogen by mass. Deduce the empirical formula of this hydrocarbon. (H = 1, C = 12)

	C	H
Step 1 the % by mass	85.7	14.3
Step 2 divide by relative atomic masses	$\frac{85.7}{12} = 7.142$	$\frac{14.3}{1} = 14.3$
Step 3 divide by the lowest number	$\frac{7.142}{7.142} = 1$	$\frac{14.3}{7.142} = 2$
Step 4 empirical formula	CH_2	

(v) Calculation involving formulae

Using molecular formula

For some compounds such as magnesium oxide, the empirical formula accurately shows the number of atoms in the compound. However, it is possible that the actual formula differs from the empirical formula. For example, the empirical formula of phosphorus(V) oxide is P_2O_5, while its actual formula is P_4O_{10}.

We call the actual formula of a compound the molecular formula.

When the empirical and molecular formulae of a compound are different, the molecular formula is always a multiple of the empirical formula.

Molecular formula = n × empirical formula, where n is 1, 2, 3, 4….

Example: A compound has the empirical formula CH_2Br. Its relative molecular mass is 188. Deduce the molecular formula of this compound. (Br = 80, C = 12, H = 1)

Step 1 empirical formula mass of $CH_2Br = 12 + (2 \times 1) + 80 = 94$

Step 2 divide the relative molecular mass by the empirical formula mass $= \frac{188}{94} = 2$

Step 3 multiply the number of atoms in the empirical formula by the number in step 2 $= 2 \times CH_2Br$

So molecular formula is $C_2H_4Br_2$.

Empirical and molecular formulae of some compounds are shown in Table 4.3.

Table 4.3 Empirical and Molecular Formulae of Some Compounds

Compound	Empirical formula	Molecular formula	Compound	Empirical formula	Molecular formula
water	H_2O	H_2O	methane	CH_4	CH_4
hydrogen peroxide	HO	H_2O_2	cyclopropane	CH_2	C_3H_6
sulphur dioxide	SO_2	SO_2	butane	C_2H_5	C_4H_{10}

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Percentage composition by mass

The formula of a compound and relative atomic masses can be used to calculate the percentage by mass of a particular element in a compound.

$$\% \text{ by mass} = \frac{\text{relative atomic mass} \times \text{number of atoms of that element in particular compound}}{\text{relative molecular (formula) mass of a compound}} \times 100$$

Example 1: Calculate the percentage by mass of iron in iron(III) oxide (Fe_2O_3). (Fe = 56, O = 16)

Step 1 Calculate the mass of an element in the compound.

(i) formula of iron(III) oxide = Fe_2O_3

(ii) relative formula mass of Fe_2O_3 = $[(2 \times 56) + (3 \times 16)] = 160$

Step 2 Calculate the percentage by mass of iron in iron(III) oxide.

(iii) $\% \text{ mass of iron} = \frac{\text{relative atomic mass of Fe} \times 2}{\text{relative formula mass of } Fe_2O_3} \times 100$

(iv) $= \frac{56 \times 2}{160} \times 100 = 70\ \%$

Example 2: Calculate the percentage by mass of nitrogen in ammonium nitrate (NH_4NO_3) fertiliser which is used by farmers to increase the yield of crops. (N = 14, H = 1, O = 16)

formula of ammonium nitrate = NH_4NO_3

relative formula mass of NH_4NO_3 = $[14 + (4 \times 1) + 14 + (3 \times 16)] = 80$

$$\% \text{ mass of nitrogen} = \frac{\text{relative atomic mass of N} \times 2}{\text{relative formula mass of } NH_4NO_3} \times 100 = \frac{14 \times 2}{80} \times 100 = 35\ \%$$

(c) Writing and Naming Formulae

It seems to be difficult to learn how to write the formulae and the name 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.

Combining capacity or valence

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. So sodium has a combining capacity of 1, calcium a combining capacity of 2 and aluminium a combining capacity of 3 in these compounds.

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(i) Fixed combining capacity 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.

(ii) 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, $CuCl_2$. Other examples include $FeCl_2$ and $FeCl_3$; Hg_2O and HgO.

Oxidation number

Oxidation number describes the combining capacity of the element and also indicates the positive and negative nature of its atoms in the compounds. 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 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.

Combining capacity and common oxidation number of some elements are shown in Table 4.4.

Table 4.4 Combining Capacity and Common Oxidation Number of Some Elements

Name	Combining capacity	Symbol with oxidation number	Name	Combining capacity	Symbol with oxidation number
sodium	1	Na^{1+}	mercury	1, 2	$Hg^{1+, 2+}$
potassium	1	K^{1+}	silver	1	Ag^{1+}
calcium	2	Ca^{2+}	carbon	2, 4	$C^{2+, 4+}$
barium	2	Ba^{2+}	sulphur	2, 4, 6	$S^{2-, 4+, 6+}$
magnesium	2	Mg^{2+}	phosphorus	3,5	$P^{3+, 5+}$
aluminium	3	Al^{3+}	oxygen	2	O^{2-}
manganese	2, 4, 7	$Mn^{2+, 4+, 7+}$	hydrogen	1	H^{1+}
zinc	2	Zn^{2+}	fluorine	1	F^{1-}
iron	2, 3	$Fe^{2+, 3+}$	chlorine	1	Cl^{1-}
tin	2, 4	$Sn^{2+, 4+}$	bromine	1	Br^{1-}
lead	2, 4	$Pb^{2+, 4+}$	iodine	1	I^{1-}
copper	1, 2	$Cu^{1+, 2+}$	nitrogen	1, 2, 3, 4, 5	$N^{3-, 1+, 2+, 3+, 4+, 5+}$

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The rules for writing and naming compounds

For convenience we will explain the formula writing and the naming of compounds in the order of : **binary compounds, acids and acid radicals, bases and basic radicals, salts and hydroxides.**

(i) Binary compounds

A binary compound is a compound which contains two elements only. Metallic binary compounds contain a metal, and non metallic binary compounds contain non-metallic elements only.

Order of symbols in a formula: 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.

positive oxidation number	**negative oxidation number**
first symbol	**second symbol**

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.

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.

$$H^{1+} + H^{1+} + S^{2-} = (H_2S)^0$$

$$\therefore \text{The formula} = H_2S$$

Other examples are:

Na and Cl	$Na^{1+} + Cl^{1-}$	$= (NaCl)^0$	$\therefore$ The formula = $NaCl$
Ca and O	$Ca^{2+} + O^{2-}$	$= (CaO)^0$	$\therefore$ The formula = CaO
Mg and Cl	$Mg^{2+} + Cl^{1-} + Cl^{1-}$	$= (MgCl_2)^0$	$\therefore$ The formula = $MgCl_2$
Al and O	$Al^{3+} + Al^{3+} + O^{2-} + O^{2-} + O^{2-}$	$= (Al_2O_3)^0$	$\therefore$ The formula = Al_2O_3

Naming binary compounds

Compounds in which the first element has a fixed oxidation number

The elements such as H^{1+}, K^{1+}, Na^{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;

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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	$BaCl_2$	Barium chloride
	CaO	Calcium oxide	AlN	Aluminium nitride

Compounds in which the first element has a variable oxidation number

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**. In those cases where there is only one atom of the first element, the use of the prefix mono is not necessary.

di, tri, tetra, etc.	Name of the first element	mono, di, tri, tetra, penta,etc.	Name of the second element ending in **-ide**

For example,

N_2O	dinitrogen monoxide (dinitrogen oxide)	NO_2	nitrogen dioxide
NO	nitrogen monoxide (nitrogen oxide)	N_2O_5	dinitrogen pentoxide
N_2O_3	dinitrogen trioxide	CCl_4	carbon tetrachloride

(2) For **naming the metallic binary compounds**, the name of the more electropositive metallic element with variable oxidation number is given first, followed by the Roman Numeral in brackets to state its oxidation number in the compound and the name of the second element ending **-ide**, is added.

Name of the first element	Oxidation number of the first element (in Roman Numeral)	Name of the second element ending in **-ide**

For example,				
	$FeCl_2$	iron(II) chloride	Cu_2O	copper(I) oxide
	$FeCl_3$	iron(III) chloride	CuO	copper(II) oxide
	PbO	lead(II) oxide	HgO	mercury(II) oxide

(ii) Acids and acid radicals

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 hydrochloric acid.

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.

In **naming the acids which have hydrogen atoms** in their molecules the name of

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acid radical **-ic** is changed to **-ide**, e.g., the acid radical of hydrochloric acid is chloride.

-ic acid, [hydrochloric acid (HCl)] $\longrightarrow$ **-ide**, [chloride Cl^-]

In **naming the acids which have oxygen atoms** in their molecules the names of the acid radicals are as follows:

-ous acid, [nitrous acid (HNO_2)] $\longrightarrow$ **-ite**, [nitrite (NO_2^-)]

-ic acid, [nitric acid (HNO_3)] $\longrightarrow$ **-ate**, [nitrate (NO_3^-)]

In **naming the acid radicals containing H**, the word hydrogen is placed before the name of the acid radical. For example, HSO_3^- = hydrogen sulphite

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 4.5.

Table 4.5 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
chloric acid	$HClO_3$	ClO_3^-	chlorate	−1	1
carbonic acid	H_2CO_3	HCO_3^- CO_3^{2-}	hydrogen carbonate carbonate	−1 −2	2
sulphurous acid	H_2SO_3	HSO_3^- SO_3^{2-}	hydrogen sulphite sulphite	−1 −2	2
sulphuric acid	H_2SO_4	HSO_4^- SO_4^{2-}	hydrogen sulphate sulphate	−1 −2	2
phosphoric acid	H_3PO_4	$H_2PO_4^-$ HPO_4^{2-} PO_4^{3-}	dihydrogen phosphate hydrogen phosphate phosphate	−1 −2 −3	3

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(iii) Bases and basic radicals

The ion formed after removal of hydroxide ions (OH^- ions) from a base is called basic radical. For example, when NaOH (base) loses OH^- ion, it forms Na^+ ion which is basic radical. The name and the oxidation number of the basic radical may be derived from the name of base as shown in Table 4.6.

Table 4.6 Bases and Basic Radicals

Name of base	Formula of base	Basic radical	Name of basic radical	Oxidation number	Number of basic radical
sodium hydroxide	$NaOH$	Na^+	sodium	+1	1
potassium hydroxide	KOH	K^+	potassium	+1	1
magnesium hydroxide	$Mg(OH)_2$	Mg^{2+}	magnesium	+2	1
calcium hydroxide	$Ca(OH)_2$	Ca^{2+}	calcium	+2	1
ammonium hydroxide	NH_4OH	NH_4^+	ammonium	+1	1

(iv) Salts

Writing the formula of a salt

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.

Metal atom or NH_4^+	Acid radical
First part	Second part

When writing the formula of a salt, the algebraic sum of the oxidation numbers must be equal to zero.

For example, to get the formula of sodium sulphate, we must combine the Na^+ with the SO_4^{2-}. In order to make the sum of the total oxidation number equal to zero, we must combine $2Na^+$ with SO_4^{2-}.

Then the sum of the oxidation numbers $= 2(+1) + (-2) = 0$

$\therefore$ The required formula $= 2Na^+ + SO_4^{2-} = Na_2SO_4$

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Other examples are:

(i)	K^+	+	NO_3^-			=	KNO_3
(ii)	Ca^{2+}	+	SO_4^{2-}			=	$CaSO_4$
(iii)	Ca^{2+}	+	HCO_3^-	+	HCO_3^-	=	$Ca(HCO_3)_2$
(iv)	NH_4^+	+	NH_4^+	+	SO_4^{2-}	=	$(NH_4)_2SO_4$

Naming salts

Naming the salt containing a metal atom with fixed 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.

$Ca(NO_3)_2$	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 begins 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.

For example:	$FeSO_4$	iron(II) sulphate
	$Fe_2(SO_4)_3$	iron(III) sulphate

Note: The Roman Numeral represents the oxidation number of the metal atom.

(v) Hydroxides

Writing the formula of hydroxide

All hydroxides include one or more - OH radicals of oxidation number − 1 in their formulae. The first part is the metal atom or the ammonium radical. The second part is the - OH radical.

Metal atom or NH_4^+	**Hydroxide radical**
First part	Second part

When writing the formula of a salt, the algebraic sum of the oxidation numbers must be equal to zero.

Naming hydroxides

Hydroxides are named in the same way as the naming of salts, but the name of acid radical is replaced by the word hydroxide. For example,

the metal atom with fixed oxidation number	$NaOH$	sodium hydroxide
	$Ca(OH)_2$	calcium hydroxide

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the metal atom with variable oxidation number	$Fe(OH)_2$	iron(II) hydroxide
	$Fe(OH)_3$	iron(III) hydroxide
the ammonium radical	NH_4OH	ammonium hydroxide

Chemistry in Society

- An important use of empirical formula calculation is in organic chemistry. Almost every day a new organic compound is either discovered or made in the laboratory. To find the formula of the new substance, a sample is analysed to obtain the mass of percentage composition of each element in the compound. From the data, the empirical formula is then worked out. The relative molecular mass and the molecular formula are also being determined.
- Sodium hydrogen carbonate (baking soda), $NaHCO_3$, is used in the manufacture of some toothpaste and as a raising agent in food production. The purity of this substance can be obtained by measuring how much carbon dioxide is given off.
- Not only chemists need to know about percentage composition, so do farmers. An important element for the growth of plants is nitrogen. If a soil is low in natural nitrogen compounds, plants do not grow as well. Therefore, a farmer can add required amount of artificial nitrogen fertiliser (urea, $CO(NH_2)_2$) to the soil.
- In the food and pharmaceutical industries, it is crucial to know the purity of the products and the formulae. Therefore, they must be labelled under the food and drugs rules and regulations.
- In some commercial products the chemical formulae are used to describe the chemical compounds.

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Review Questions

(1) Write the empirical formulae for (a) hydrazine, N_2H_4 (b) octane, C_8H_{18} (c) benzene, C_6H_6 and (d) ammonia, NH_3.

(2) The composition by mass of a hydrocarbon is 10 % hydrogen and 90 % carbon. Deduce the empirical formula of this hydrocarbon. (C = 12, H = 1)

(3) Write the formula of each of the following compounds:
magnesium sulphate, potassium carbonate, lead(II) chloride, zinc oxide, ammonium sulphate, aluminium chloride, sulphur trioxide, sodium bromide

Key Terms

- **Empirical formula** shows the simplest whole number ratio of atoms in a compound.
- **Molecular formula** shows the total number of atoms of each element present in one molecule or one formula unit of the compound.
- **Oxidation number** describes the combining capacity of the element and also indicates the positive and negative nature of its atoms in the compounds. The oxidation number is related to, but not identical with valence or combining capacity.

4.3 CHEMICAL EQUATIONS

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. A balanced chemical equation helps us to calculate the right amount of reactant to use in a reaction.

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, and
(4) the weight relationship of the reactants and the products.

There are two types of chemical equations: **equation in words** and **equation in symbols**.

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.

Equation in words: carbon + oxygen $\longrightarrow$ carbon dioxide

Equation in symbols: $C + O_2 \longrightarrow CO_2$

(i) Steps in writing chemical equations

Let us take the burning of magnesium as an example.

Step 1	Write the word equation for the reaction.	magnesium + oxygen $\longrightarrow$ magnesium oxide
Step 2	Write in symbols and the formulae of the reactants and the product under the respective names.	$Mg + O_2 \longrightarrow MgO$
Step 3	Balance the equation so that the number of atoms of each element is equal on both sides of the equation.	$2Mg + O_2 \longrightarrow 2MgO$

In Step 3 a complete balanced chemical equation has to abide the **Law of Conservation of Mass**. That is, the total mass of the reactant(s) is equal to the total mass of the product(s). 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.

(ii) The physical states of the reactants and products

A complete balanced chemical equation must also show the physical states of the reactants and products, whether they are solid, liquid, gaseous, aqueous or in solution. The abbreviations of physical states are written after the corresponding symbols and formulae.

The following is an example of a complete balanced equation including the physical states:

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$$2Na(s) + 2H_2O(l) \longrightarrow 2NaOH(aq) + H_2(g)$$

Note: 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

For example,

$AgNO_3(aq) + NaCl(aq) \longrightarrow AgCl(s)\downarrow + NaNO_3(aq)$ (precipitation reaction)

$2Mg(s) + O_2(g) \xrightarrow{\Delta} 2MgO(s)$ (combination reaction)

$S(s) + O_2(g) \xrightarrow{\Delta} SO_2(g)$ (combustion)

$NH_4Cl(s) \rightleftharpoons NH_3(g)\uparrow + HCl(g)\uparrow$ (reversible reaction)

(iii) Writing ionic equations

Most ionic compounds are soluble in water. They exist as ions in aqueous solution. An ionic equation is a simplified chemical equation that shows the reactions involving ions in solution. Let us now see how an ionic equation is written.

When aqueous sodium chloride is added to aqueous silver nitrate, a white precipitate of silver chloride is formed. This reaction can be represented by the following balanced chemical equation:

silver nitrate + sodium chloride $\longrightarrow$ silver chloride + sodium nitrate

$$AgNO_3(aq) + NaCl(aq) \longrightarrow AgCl(s)\downarrow + NaNO_3(aq)$$

Since $AgNO_3$, NaCl and $NaNO_3$ are soluble in water they exist as ions in aqueous solution.

The chemical equation in terms of ions can be written as:

$$Ag^+(aq) + NO_3^-(aq) + Na^+(aq) + Cl^-(aq) \longrightarrow AgCl(s)\downarrow + Na^+(aq) + NO_3^-(aq)$$

The Na^+ ions and NO_3^- ions are still ions in solution at the end of the reaction. They have not taken part in the chemical reaction. Such ions are called 'spectator ions'.

Since only Ag^+ ions and Cl^- ions have reacted, the equation for the reaction can therefore be simplified as shown below.

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

This is the ionic equation for the reaction between aqueous silver nitrate and aqueous sodium chloride.

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Review Questions

(1) Write the balanced chemical equations for the following reactions:
   (a) sodium + chlorine $\longrightarrow$ sodium chloride
   (b) sodium + water $\longrightarrow$ sodium hydroxide + hydrogen
   (c) zinc + sulphuric acid $\longrightarrow$ zinc sulphate + hydrogen

(2) When zinc is added to copper(II) sulphate solution, copper and zinc sulphate are formed. Write the balanced ionic equation for this reaction.

Key Terms

- The substances that take part in the reaction are called **reactants**.
- The substances that are produced in the reaction are called **products**.
- **Law of Conservation of Mass** states that the total mass of the reactant(s) is equal to the total mass of the product(s).

4.4 THE MOLE CONCEPT

The relative atomic masses of elements as well as the relative molecular masses of ions or compounds expressed in amu, are not feasible to be used in the weighable scale, especially in grams or other weighable masses. The relative atomic and molecular masses are also composed of microscopic particles and therefore impossible to count individually. The quantitative use of balanced chemical equations so as to calculate the reactants and products need for translating the 'amu' with weighable gram or kilogram, i.e., in the big scale. The gram or kilogram is the measure of countable unit whereas amu is not countable by weighing. For instance, C-12 which has 12 amu becomes 12 g; Mg-24 amu becomes 24 g, and He-4 amu becomes 4 g. It is clear that atomic mass unit (amu) and gram (g) or kilogram (kg) are interrelated. Numerically, the relative atomic and molecular masses are pure numbers.

The chemists' counting unit is named the mole, and it is defined as equal to the number of atoms in exactly 12 g of ^{12}C. This number of particles is called **Avogadro's constant**, which is equal to 6.02×10^{23}.

(a) The Mole and the Avogadro's Constant

We may refer to the mass of a mole of substance as its molar mass (abbreviation M). The unit of molar mass is g mol^{-1}. It is the relative atomic, molecular or formula mass of the substance in grams.

The number of atoms in a mole of atoms is very large: 6.02×10^{23} atoms. This number is called the Avogadro's constant or Avogadro's number (L).

One mole of a substance is the amount of substance that has the same number of particles (atoms, molecules, etc.) as there are atoms in exactly 12 g of ^{12}C.

The Avogadro's constant applies to atoms, molecules, ions and electrons. So in 1 mole of sodium there are 6.02×10^{23} sodium atoms, and in 1 mole of sodium chloride (NaCl) there are 6.02×10^{23} sodium ions and 6.02×10^{23} chloride ions. A mole of chlorine molecules, Cl_2, contains 6.02×10^{23} chlorine molecules but twice as many chlorine atoms, as there are two chlorine atoms in every chlorine molecule.

Amedeo Avogadro (1776-1856) was an Italian scientist who first deduced that equal volumes of gases contain equal number of molecules under the same conditions of temperature and pressure.

(b) Molar Volume of the Gas

One mole of any gas has a volume of 24 dm^3 or 24,000 cm^3 at room temperature and pressure (r.t.p.). This volume is called the **molar volume of a gas**.

Room temperature and pressure are often taken as conditions of 25 °C and 1 atmosphere.

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To convert volumes of gases into moles and moles of gases into volumes, the following relationship is used.

$$\text{number of mole of a gas} = \frac{\text{volume of the gas (in dm}^3\text{)}}{\text{molar volume of the gas at r. t. p.}}$$

$$\text{number of mole of a gas} = \frac{\text{volume of the gas (in dm}^3\text{)}}{24\ \text{dm}^3\,\text{mol}^{-1}\ \text{at r. t. p.}}$$

Note: One mole of every gas occupies 22.4 dm^3 at STP (standard temperature, 0 °C or 273 K and standard pressure, 760 mmHg or 1 atmosphere).

Example: Calculate the volume of 0.5 mol of carbon dioxide at room temperature and pressure (r.t.p.).

$$\text{number of moles of } CO_2 = \frac{\text{volume dm}^3 \text{ of } CO_2}{24\ \text{dm}^3\,\text{mol}^{-1}\ \text{at r. t. p.}}$$

$$\text{volume of } CO_2 \text{ at r.t.p.} = \text{number of moles of } CO_2 \times 24\ \text{dm}^3\,\text{mol}^{-1}$$

$$= 0.5 \times 24 = 12\ \text{dm}^3 = \mathbf{12{,}000\ cm^3}$$

(c) Moles and Mass

The Système International (SI) unit for mass is the kilogram. But this is a rather large mass to use for general laboratory work in chemistry. So chemists prefer to use the relative molecular mass or formula mass in grams (1000 g = 1 kg). You can find the number of moles of a substance by using the mass of substance and the relative atomic mass or relative molecular mass or molar mass.

$$\text{number of moles (mol)} = \frac{\text{mass of a substance in gram (g)}}{\text{molar mass (g mol}^{-1}\text{)}}$$

To find the mass of a substance present in a given number of moles, you need to rearrange the equation:

$$\text{mass of substance (g)} = \text{number of moles (mol)} \times \text{molar mass (g mol}^{-1}\text{)}$$

Example 1: How many moles of sodium chloride are present in 117.0 g of sodium chloride, NaCl? (Na = 23, Cl = 35.5)

$$\text{molar mass of NaCl} = 23.0 + 35.5 = 58.5\ \text{g mol}^{-1}$$

$$\text{number of moles} = \frac{\text{mass}}{\text{molar mass}} = \frac{117.0\ \text{g}}{58.5\ \text{g mol}^{-1}} = \mathbf{2.0\ mol}$$

Example 2: What mass of sodium hydroxide, NaOH, is present in 0.25 mol of sodium hydroxide? (H = 1, Na = 23, O = 16)

$$\text{molar mass of NaOH} = 23 + 16 + 1 = 40\ \text{g mol}^{-1}$$

$$\text{mass} = \text{number of moles} \times \text{molar mass}$$

$$= 0.25\ \text{mol} \times 40\ \text{g mol}^{-1} = \mathbf{10\ g\ NaOH}$$

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Example 3: Calculate the total number of molecules in 7.10 g of chlorine molecule (Cl_2).
(Avogadro's constant = 6.02×10^{23} mol^{-1}; Cl = 35.5)

molar mass of $Cl_2 = 2 \times 35.5 = 71$ g mol^{-1}

$$\text{number of moles of chlorine} = \frac{\text{mass of chlorine molecule}}{\text{molar mass of chlorine}} = \frac{7.1\ \text{g}}{71\ \text{g mol}^{-1}} = 0.1\ \text{mol}$$

$$\begin{aligned}\text{number of molecules of } Cl_2 &= \text{number of moles of chlorine} \times 6.02 \times 10^{23} \text{ molecules mol}^{-1}\\ &= 0.1 \text{ mol} \times 6.02 \times 10^{23} \text{ molecules mol}^{-1}\\ &= \mathbf{6.02 \times 10^{22} \text{ molecules}}\end{aligned}$$

(d) Mole Calculations

(i) Calculations from equations (Reacting masses)

Example 1: Calculate the mass of water produced from the complete combustion of 0.25 mol of methane.

$$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(l)$$

Step1 Write the balanced equation.

$$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(l)$$

Step 2 From the equation, find the ratio of the number of moles of H_2O to the number of moles of CH_4.

$$\frac{\text{number of moles of } H_2O \text{ produced}}{\text{number of moles of } CH_4 \text{ reacted}} = \frac{2}{1}$$

Step 3 Use the ratio to find the number of moles of H_2O produced when 0.25 mole of CH_4 is burnt.

$$\begin{aligned}\text{number of moles of } H_2O &= 2 \times \text{number of moles of } CH_4\\ &= 2 \times 0.25 \text{ mol} = 0.5 \text{ mol}\end{aligned}$$

Step 4 Multiply the number of moles by the molar mass of H_2O to obtain the mass of H_2O in grams.

$$\text{molar mass of } H_2O = (2 \times 1) + 16 = 18 \text{ g mol}^{-1}$$

$$\begin{aligned}\text{mass of } H_2O \text{ in grams} &= \text{number of moles} \times \text{molar mass of } H_2O\\ &= 0.5 \text{ mol} \times 18 \text{ g mol}^{-1} = \mathbf{9\ g}\end{aligned}$$

(ii) Calculations from equations (Reacting masses and volumes)

Example 2: Magnesium reacts with hydrochloric acid according to the equation:

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

Calculate the volume of hydrogen gas, measured at room conditions, produced from the reaction of 14.6 g of hydrochloric acid.

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Step 1 Change the mass of HCl into moles.

molar mass of HCl $= 1 + 35.5 = 36.5$ g mol^{-1}

$$\text{number of moles of HCl} = \frac{\text{mass of HCl (g)}}{\text{molar mass of HCl (g mol}^{-1})} = \frac{14.6}{36.5} = 0.4 \text{ mol}$$

Step 2 Write the chemical equation.

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

Step 3 From the equation, find the ratio of the number of moles of H_2 to the number of moles of HCl.

$$\frac{\text{number of moles of } H_2}{\text{number of moles of HCl}} = \frac{1}{2}$$

Step 4 Use the ratio to find the number of moles of H_2 produced when 0.4 mole of HCl reacts.

$$\text{number of moles of } H_2 = \frac{1}{2} \times \text{number of moles of HCl} = \frac{1}{2} \times 0.4 \text{ mol} = 0.2 \text{ mol}$$

Step 5 Multiply the number of moles of H_2 gas by the molar gas volume. This gives the volume of H_2 gas produced.

volume of H_2 gas = number of moles × molar gas volume = 0.2×24 dm^3 = **4.8 dm^3**

Chemistry in Society

- **Medicine**: In order to make drug from its ingredients, someone has to figure out how much of each ingredient is needed to react together to make the final drug. That would have involved using the concept of moles.
- **Plastic**: Some plastics are made from other chemicals, someone has to figure out how much of each ingredient is needed to use, and that would have involved moles.
- **Combustion**: You need to use mole in combustion to know how much air is needed, how much exhaust would be produced, as well as how much heat is evolved.
- **Batteries**: Chemicals in batteries react to produce electricity. People have to figure out how much of each type of chemical is needed to put together in a battery to make its function properly. They would also need to know how much the amount of moles of each reactant is needed.

Review Questions

(1) Calculate the amount of moles in 10.7 g of sulphur atoms. (S = 32)

(2) What is the mass of 0.20 mol of carbon dioxide, CO_2? (C = 12, O = 16)

(3) You have a 56 g sample of iron(II) sulphide, FeS.

(a) How many moles of FeS are there in the sample?

(b) How many molecules of FeS are there in the sample?

(Fe = 56, S = 32, Avogadro's constant = 6.02×10^{23} mol^{-1}).

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(4) How many moles are present in the following volumes of gases at r.t.p.?
(a) 1.2 dm^3 of sulphur dioxide (SO_2) (b) 0.24 dm^3 of methane (CH_4)
(c) 120 cm^3 of carbon dioxide (CO_2)

Key Terms

- **One mole** of a substance is the amount of substance that has the same number of particles (atoms, molecules, etc.) as there are atoms in exactly 12 g of ^{12}C.
- The mass of one mole of a substance is called the **molar mass**.
- Equal volumes of gases contain equal number of molecules at the same temperature and pressure.
- The Avogadro's constant (**Avogadro's number** = 6.02×10^{23}) is the number of entities or a stated type of particles (atoms, ions or molecules) in a mole of those substances.
- One mole of any gas has a volume of 24 dm^3 or 24,000 cm^3 at room temperature and pressure (r.t.p.). This volume is called the **molar volume of a gas**.
- One mole of every gas occupies 22.4 dm^3 at **STP** (standard temperature, 0 °C or 273 K and standard pressure, 760 mmHg or 1 atmosphere).

EXERCISES

1. Match each of the items given in List A with the appropriate item given in List B.

List A	List B
(a) Number of acid radicals in H_2SO_4	(i) 24 dm^3 at r.t.p.
(b) The mass of a compound of giant structure	(ii) MgO
(c) The formula of magnesium oxide	(iii) molar mass
(d) Molar volume of gas	(iv) formula mass
(e) The mass of a mole of substance	(v) 2

2. Calculate the empirical formula of each of the following compounds:
(a) A compound in which 3 g of carbon combines with 4 g of oxygen.
(b) Iron oxide in which the weight of iron is 77.7 % and that of oxygen is 22.3 %.
(c) Water in which hydrogen and oxygen combine in the proportion of 1:8 by weight.
(Fe = 56, C = 12, O = 16, H = 1)

3. The empirical formulae and relative molecular masses of three compounds, A, B and C are shown as follows. Calculate the molecular formula of each of these compounds.
(C = 12, Cl = 35.5, H = 1)

Compound	Empirical formula	Relative molecular mass
A	C_3H_5	82
B	CCl_3	237
C	CH_2	112

4. Hydrocarbons are compounds of carbon and hydrogen only. Hydrocarbon Z is composed of 80 % carbon and 20 % hydrogen.

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(a) Calculate the empirical formula of hydrocarbon Z. (C = 12, H = 1)
(b) The relative molecular mass of hydrocarbon Z is 30. Deduce the molecular formula of this hydrocarbon.

5. Vinegar, which is used in our homes, is a dilute form of acetic acid. A sample of acetic acid has the following percentage composition: 39.9 % carbon, 6.7 % hydrogen and 53.4 % oxygen.
(a) Determine the empirical formula of acetic acid.
(b) Determine the molecular formula of acetic acid if the molecular mass of acetic acid is 60 amu. (C = 12, H = 1, O = 16)

6. Hydrogen peroxide decomposes according to the equation:

$$2H_2O_2(l) \longrightarrow 2H_2O(l) + O_2(g)$$

Calculate the volume of oxygen gas produced at r.t.p. when 1.7 g of H_2O_2 is decomposed. (H = 1, O = 16)

7. Tin(IV) oxide is reduced to tin by heating with carbon. Carbon monoxide is also formed.

$$SnO_2(s) + 2C(s) \xrightarrow{\Delta} Sn(s) + 2CO(g)$$

Calculate the mass of carbon that exactly reacts with 14 g of tin(IV) oxide.
(C = 12, O = 16, Sn = 119)

8. A conical flask contains 68.4 g of octane (C_8H_{18}). How many molecules of octane are there in the flask? (Avogadro's constant = 6.02×10^{23} mol^{-1}; C = 12, H = 1).

9. Calculate the number of atoms in 4 g of bromine molecule (Br_2).
(Avogadro's constant = 6.02×10^{23} mol^{-1}; Br = 80).

10. Solid sodium carbonate reacts with aqueous hydrochloric acid to form aqueous sodium chloride, carbon dioxide and water.

$$Na_2CO_3 + 2HCl \longrightarrow 2NaCl + CO_2 + H_2O$$

(a) Rewrite this equation including state symbols.
(b) Calculate the number of moles of hydrochloric acid required to react exactly with 4.15 g of sodium carbonate. (C = 12, Na = 23, O = 16)

11. Identify the spectator ions and write the ionic equations for the following reactions:
(a) $Na_2SO_4(aq) + Ba(NO_3)_2(aq) \longrightarrow BaSO_4(s) + 2NaNO_3(aq)$
(b) $CaCO_3(s) + 2HCl(aq) \longrightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$
(c) $(NH_4)_2SO_4(aq) + 2NaOH(aq) \longrightarrow Na_2SO_4(aq) + 2H_2O(l) + 2NH_3(g)$

12. Calculate the amount of substance in moles in each of the following:
(a) 64.2 g of sulphur molecules (S_8)
(b) 60.45 g of anhydrous iron(III) nitrate, $Fe(NO_3)_3$. (Fe = 56, N = 14, O = 16, S = 32)

13. Calculate the mass in grams of the following:
(a) 0.050 moles of sodium carbonate, Na_2CO_3
(b) 5.00 moles of iron(II) hydroxide, $Fe(OH)_2$
(C = 12, Fe = 56, H = 1, O = 16, Na = 23)

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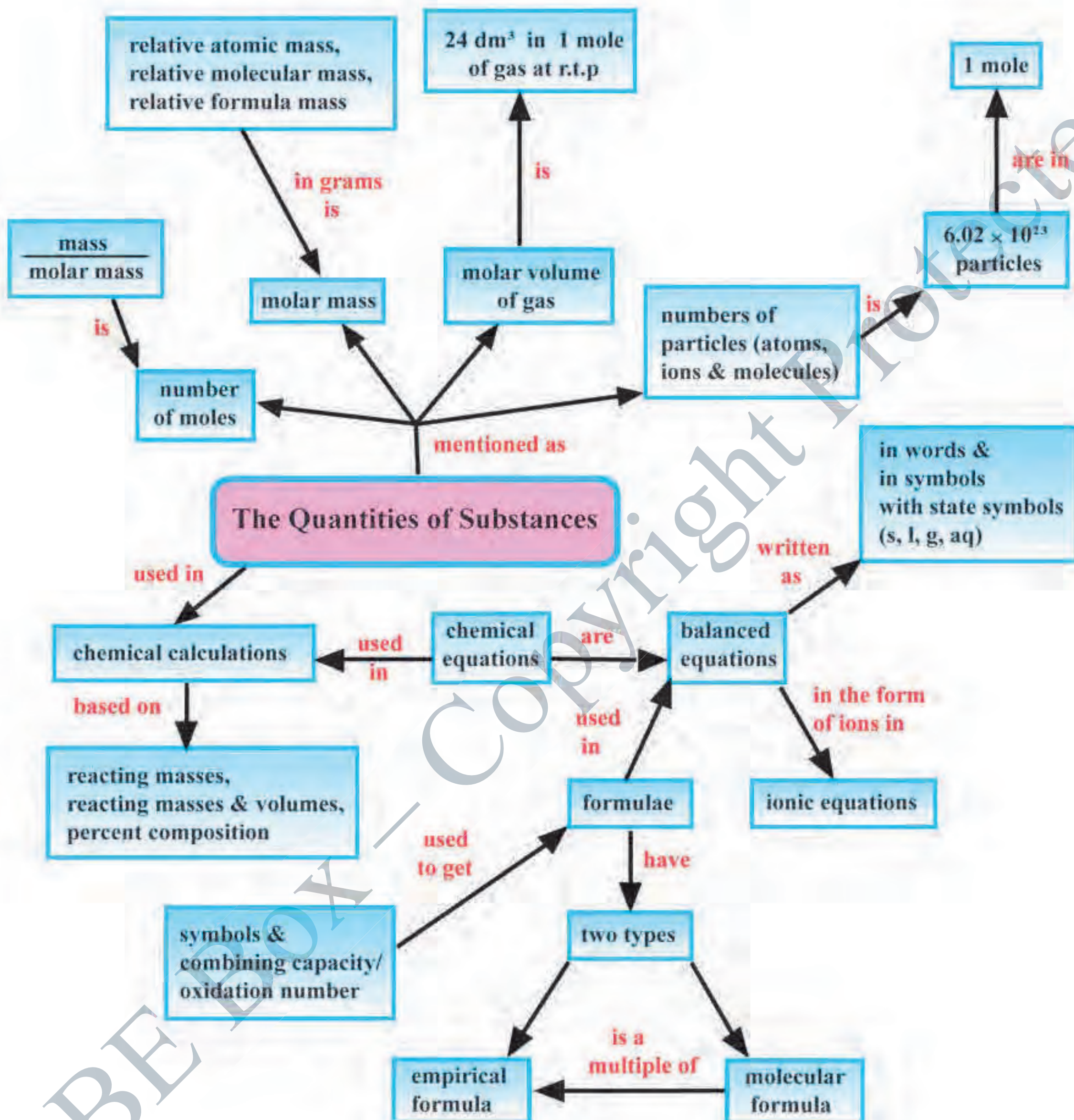
CHAPTER REVIEW
(Concept Map)



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CHAPTER 5 NON-METALS: OXYGEN, CARBON AND HALOGENS

About 50 % of the mass of the Earth's crust consists of oxygen. In combination with carbon, oxygen, hydrogen and nitrogen occur in a large part of plants and animals. Living organisms are mostly made of non-metals. Roughly 96 % of the mass of the human body is made up of just four elements: oxygen, carbon, hydrogen and nitrogen. In this chapter, properties and uses of three non-metals such as oxygen, carbon and halogens (fluorine, chlorine, bromine and iodine) are studied.

You have already learned in Chapter 3 that oxygen with symbol O is the eighth element of the Periodic Table found in the group VI and period 2. The atomic number of oxygen is 8; the mass number is 16. Therefore, the symbol for oxygen atom is $^{16}_{8}O$. Gaseous oxygen molecule is written as O_2.

Carbon is a chemical element with symbol C and atomic number 6, the mass number 12 in group IV and period 2 of the Periodic Table. The symbol is $^{12}_{6}C$. Today, C-12 (exactly) is the standard representative definition of atomic mass unit of all the elements in the Periodic Table.

Halogens are in group VII in the Periodic Table. It consists of five elements: fluorine, chlorine, bromine, iodine and the radioactive element astatine. Halogens are the most reactive non-metals. They react with most metals to form salts. 'Halogen' means **salt-former** in Greek. The molecular formulae are written as F_2, Cl_2, Br_2 and I_2.

Learning Outcomes

After completing this chapter, students will be able to:
- describe the properties and behaviours of oxygen and oxides;
- classify the main types of oxides based on their properties;
- describe the properties and behaviours of carbon;
- explain the allotropy and allotropes of carbon;
- distinguish and compare the properties and behaviours of halogens and halides;
- recognise the role of oxygen, oxides, carbon and halogens in daily life.


Oxygen cylinder


Carbon (Diamond)


Chlorine $Cl_2(g)$


Bromine $Br_2(l)$


Iodine $I_2(s)$

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5.1 OXYGEN

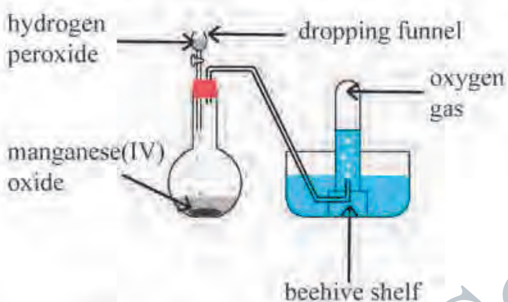
Oxygen is the most abundant element in the Earth's crust. The free element accounts for 21 % of the volume of the atmosphere. Oxygen in the combined state exists in water, sand or silica, silicates and rocks.

(a) Preparation of Oxygen

Oxygen preparation can be demonstrated in a number of ways in the classroom or in the laboratory. Chlorates are principally toxic by injection and inhalation. Therefore, potassium chlorate should not be used for the preparation of oxygen in the laboratory.

Activity (1): Preparation of oxygen from hydrogen peroxide

In the laboratory or in the classroom, oxygen gas can be prepared by using an environmentally friendly liquid hydrogen peroxide of appropriate strength. In this reaction manganese(IV) oxide is used as a catalyst to speed up the chemical reaction. Oxygen gas is collected by the downward displacement of water.



$$\text{hydrogen peroxide} \xrightarrow{MnO_2} \text{water} + \text{oxygen}$$

$$2H_2O_2(aq) \xrightarrow{MnO_2} 2H_2O(l) + O_2(g)$$

(b) Properties of Oxygen

Physical properties

Oxygen is a colourless gas, without taste or smell. Oxygen is only slightly soluble in water and has about the same relative vapour density as air. Oxygen will not burn, however, it supports combustion. It rekindles any glowing thing (splinter) with a bright flame.

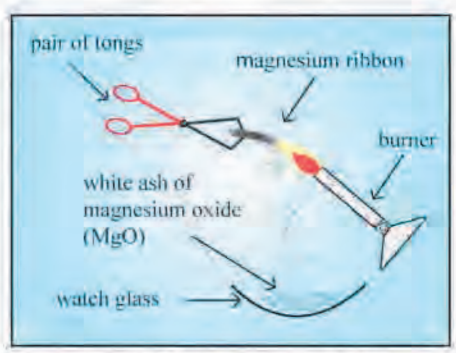
Chemical properties

(i) Action with metals

Metals such as magnesium, iron, copper and zinc react with oxygen forming oxides. A piece of clean magnesium ribbon continues burning in oxygen from air with a dazzling white flame, leaving a white powder as residue. This residue is magnesium oxide.

$$\text{magnesium} + \text{oxygen} \xrightarrow{\Delta} \text{magnesium oxide}$$

$$2Mg(s) + O_2(g) \xrightarrow{\Delta} 2MgO(s)$$



Burning of magnesium ribbon

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Iron slowly becomes oxidised in the presence of air and water, to form hydrated iron(III) oxide, i.e., iron has become rusted.

iron + oxygen + water $\longrightarrow$ hydrated iron(III) oxide (rust)

$$4Fe(s) + 3O_2(g) + 2nH_2O(l) \longrightarrow 2[Fe_2O_3.nH_2O](s)$$

Most useful metals or sheets become covered with thin films of metal oxide.

(ii) Action with non-metals

Non-metals such as phosphorus, sulphur and carbon react with oxygen forming oxides. A small piece of phosphorus (only red phosphorus may be used) burns in oxygen giving off white fumes which consist of oxides of phosphorus.


Burning red phosphorus in air

phosphorus + oxygen $\longrightarrow$ phosphorus(V) oxide

$$4P(s) + 5O_2(g) \longrightarrow P_4O_{10}(s)$$

phosphorus + oxygen $\longrightarrow$ phosphorus(III) oxide

$$4P(s) + 3O_2(g) \longrightarrow P_4O_6(s)$$

(c) Oxides

An oxide is a compound containing oxygen and another element. There are six main types of oxides.

(i) acidic oxides (ii) basic oxides (iii) amphoteric oxides
(iv) neutral oxides (v) peroxides (vi) compound oxides

(i) Acidic oxides

An acidic oxide is an oxide of non-metal. It reacts with basic oxides to give salts only.

carbon dioxide + sodium oxide $\longrightarrow$ sodium carbonate

$$CO_2(g) + Na_2O(s) \longrightarrow Na_2CO_3(s)$$

Acidic oxides react with alkali solutions to give salts and water.

Test for carbon dioxide

carbon dioxide + calcium hydroxide $\longrightarrow$ calcium carbonate + water

$$CO_2(g) + Ca(OH)_2(aq) \longrightarrow CaCO_3(s) + H_2O(l)$$

Some acidic oxides are soluble in water but some are not. CO_2, SO_2, SO_3 and P_4O_{10} are soluble oxides while SiO_2 is an insoluble oxide. Soluble acidic oxides dissolve in water to form acidic solutions. These solutions turn blue litmus red.

sulphur trioxide + water $\longrightarrow$ sulphuric acid

$$SO_3(g) + H_2O(l) \longrightarrow H_2SO_4(aq)$$

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(ii) Basic oxides

A basic oxide is an oxide of metal. A basic oxide reacts with an acid to produce a salt and water only. It neutralises acids.

sodium oxide + hydrochloric acid (dil.) $\longrightarrow$ sodium chloride + water

$Na_2O(s) + 2HCl(aq) \longrightarrow 2NaCl(aq) + H_2O(l)$

Some basic oxides are soluble in water but some are not. Na_2O and K_2O are soluble oxides while MgO, CuO and Ag_2O are not. Some basic oxides react with water forming hydroxide solutions (alkalis). These solutions turn red litmus blue.

sodium oxide + water $\longrightarrow$ sodium hydroxide

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

(iii) Amphoteric oxides

An amphoteric oxide is a metallic oxide which possesses both basic and acidic properties. It reacts with both acids and alkalis to form salt and water. (e.g., ZnO, Al_2O_3, PbO)

zinc oxide + sulphuric acid (dil.) $\longrightarrow$ zinc sulphate + water

$ZnO(s) + H_2SO_4(aq) \longrightarrow ZnSO_4(aq) + H_2O(l)$

zinc oxide + sodium hydroxide $\longrightarrow$ sodium zincate + water

$ZnO(s) + 2NaOH(aq) \longrightarrow Na_2ZnO_2(aq) + H_2O(l)$

(iv) Neutral oxides

The neutral oxide is an oxide which shows neither basic nor acidic character. (e.g., CO, N_2O)

(v) Peroxides

Those oxides that react with an acid to give salt and hydrogen peroxide are called peroxides. (e.g., BaO_2, Na_2O_2)

barium peroxide + sulphuric acid (dil.) $\longrightarrow$ barium sulphate + hydrogen peroxide

$BaO_2(s) + H_2SO_4(aq) \longrightarrow BaSO_4(s) + H_2O_2(aq)$

PbO_2, MnO_2 and NO_2 are not peroxides. They do not give hydrogen peroxides on reaction with acids.

(vi) Compound oxides

A compound oxide is an oxide, formed by the combination of two different oxides of the same element. (e.g. Pb_3O_4, Mn_3O_4, Fe_3O_4)

Pb_3O_4 is a compound oxide consisting of lead(II) and lead(IV) oxides. It can be written as di-lead(II) lead(IV) oxide or red lead oxide, ($2PbO.PbO_2$). When red lead oxide reacts with dilute nitric acid, lead(II) oxide only reacts with the acid to form lead(II) nitrate.

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red lead oxide + nitric acid (dil.) ⟶ lead(II) nitrate + water + lead(IV) oxide

$$Pb_3O_4(s) + 4HNO_3(aq) \longrightarrow 2Pb(NO_3)_2(aq) + 2H_2O(l) + PbO_2(s)$$

Chemistry in Society

- Man and animals use oxygen from the air in respiration. Oxygen is used to help patients with breathing difficulties. Mountaineers and under water divers need oxygen cylinders with them. In aquatic habitat, organisms use oxygen dissolved in water.
- In living organisms, the oxygen intake is used for the breakdown of the glucose molecules to produce energy.
- Oxygen is used in steel work, and oxyacetylene gas mixture is used for cutting and welding metals.
- Liquid oxygen is carried on rockets to support the fuel burns. Oxygen is essential in combustion processes such as the burning of fuels.

Usefulness of some oxides in society is described in the following Table:

Oxides	Uses	Oxides	Uses
CO_2	fire extinguishers	SiO_2	manufacture of glass
SO_2	bleaching agent	Pb_3O_4	pigment used in paints
MgO	laxative	ZnO	skin conditioners, cosmetics
CaO	cement production	H_2O_2	antiseptic, hair dyes and toothpaste

Review Questions

(1) The mountaineers and under water divers need to carry oxygen cylinders. Why?

(2) Why manganese(IV) oxide is used as a catalyst for the preparation of oxygen in the laboratory?

(3) Identify the class of oxides to which each of the following belongs:

(a) carbon monoxide (b) red lead oxide (c) sulphur dioxide

(d) sodium peroxide (e) copper(II) oxide (f) lead(II) oxide

Key Terms

- An **acidic oxide** is a non-metallic oxide which reacts with basic oxide to produce salt.
- A **basic oxide** is a metallic oxide which reacts with acid to form salt and water.
- An **amphoteric oxide** is a metallic oxide which possesses both basic and acidic properties.
- A **neutral oxide** does not react with either acids or bases.
- A **peroxide** reacts with an acid to produce salt and hydrogen peroxide.
- A **compound oxide** is the combination of two different oxides of the same element.

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5.2 CARBON

Carbon is found in nature as diamond and graphite. Fullerene and graphene are synthetic carbon. Carbon can also be found as compounds combining with other elements in petroleum, coal, natural gas, limestone, carbon dioxide and sugar ($C_{12}H_{22}O_{11}$), etc. In addition, all living things have carbon containing compounds such as carbohydrates, fats, proteins and nucleic acids, etc.

(a) Allotropy and Allotropes of Carbon

If an element, can exist more than one form, in the same physical state, it is said to exhibit allotropy or polymorphism. The different forms of an element in the same physical state that possess different physical properties are known as allotropes of that element. They may have different chemical properties. For example,

Diamond, graphite, fullerene and graphene are allotropes of carbon.
Oxygen and ozone are allotropes of oxygen.
Rhombic sulphur and monoclinic sulphur are allotropes of sulphur.

Diamond

In diamond, each carbon atom is surrounded by four other carbon atoms (Figure 5.1). It has a giant structure. It contains millions of carbon atoms in a three dimensional network of strong carbon-carbon covalent bonds. Therefore, it is very hard and has a very high melting point (3550 °C). Diamond is the hardest among all naturally occurring substances. It is transparent and shines in presence of light.

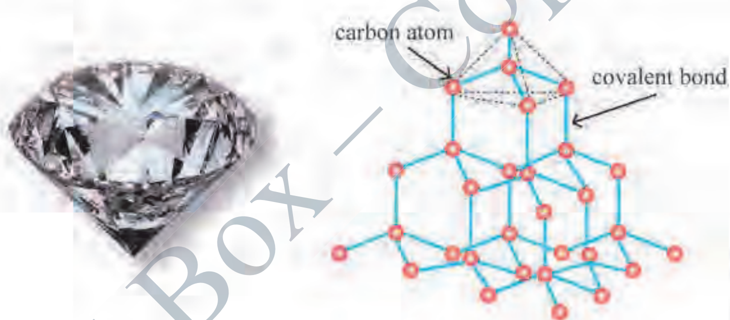

Figure 5.1 Diamond and Its Structure

Graphite

In graphite, each carbon atom is surrounded by three other carbon atoms in the same plane, and therefore layers of hexagons are obtained (Figure 5.2). 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.

Due to its layered structure, graphite is soft and has soapy touch. As the layers are bonded through weak forces known as van der Waals forces, it can act as a lubricant.

Due to the presence of free electrons, it is a good conductor of electricity and heat. The melting point of graphite is 3700 °C.

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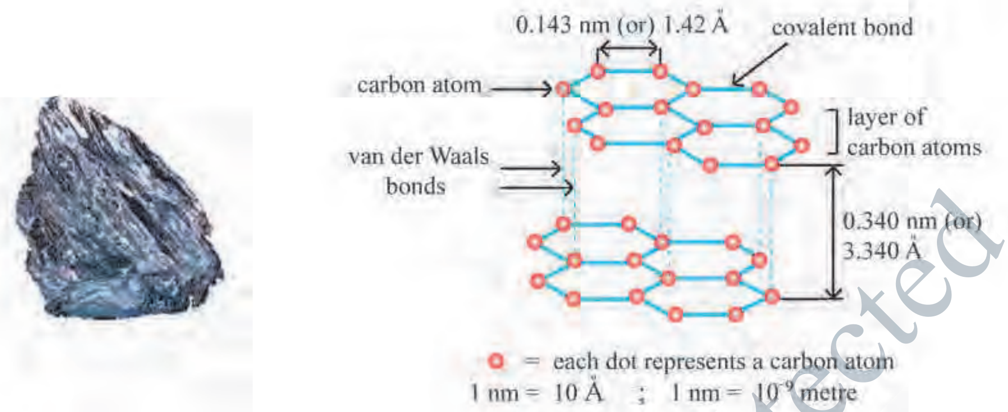

Figure 5.2 Graphite and Its Structure

Fullerene

Fullerene (C_{60}) is an allotrope of carbon in the form of a hollow sphere, ellipsoid, tube and many other shapes and sizes. Spherical fullerenes, also referred to as Bucky balls, resemble the balls used in association football. Cylindrical fullerenes are also called carbon nanotubes (Figure 5.3). Fullerenes are stable, but not totally unreactive. Fullerenes cannot be found in nature. They are the synthetic allotropes of carbon. Fullerenes have lower melting and boiling points than diamond and graphite.

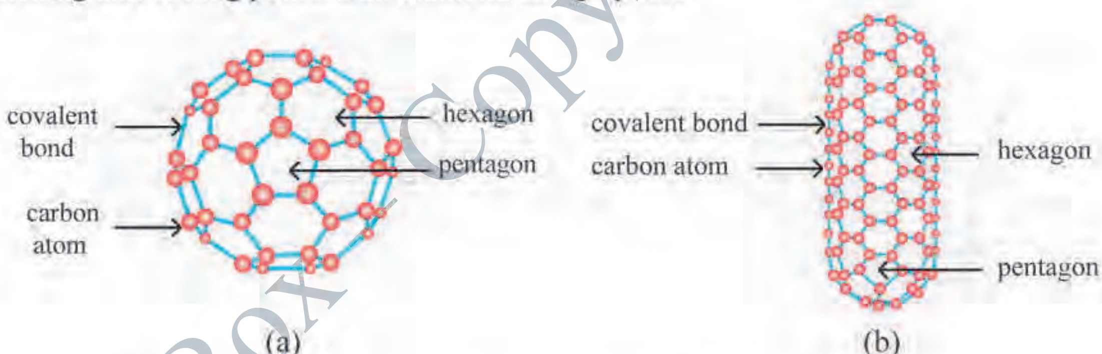

Figure 5.3 Fullerenes (a) Bucky Ball (spherical fullerene)
(b) Carbon Nanotube (cylindrical fullerene)

Graphene

Graphene is an allotrope of carbon. Graphene cannot be found in nature. It is a single layer (monolayer) of graphite (Figure 5.4). It is tightly bound in a hexagonal ring structure. Its crystalline structure is two-dimensional.

Graphene has many properties. In proportion to its thickness, it is about 100 times stronger than the strongest steel. Graphene is a transparent and flexible conductor so that it is widely used for various material/device applications, including solar cells, light-emitting diodes (LED), touch panels and smart windows or phones.

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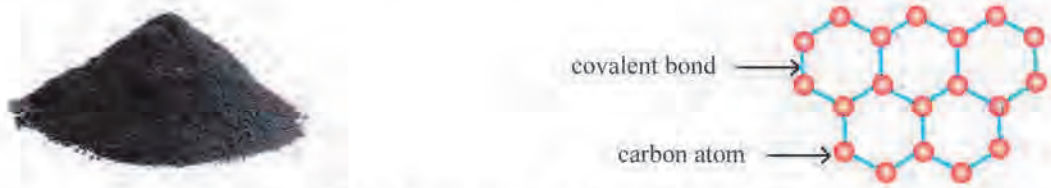

Figure 5.4 Graphene and Its Structure

(b) Other Forms of Carbon

Charcoal, coal, coke and carbon black (soot) 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 is made by heating wood in the absence of air. It has a porous structure and has many small holes.

Coal is found in nature. In coal the element carbon is mixed with compounds of other elements. Coal is a black heavy solid.

Coke is formed by heating coal in the absence of air. Coke is also a black heavy solid. It is almost pure carbon.

Carbon black (soot) is a black powder. When kerosene is burnt in a limited amount of air, hydrogen from kerosene combines with oxygen from the air and the carbon is left as carbon black.

(c) Chemical Properties of Carbon

(i) When carbon burns in excess air or oxygen, carbon dioxide is formed. Carbon monoxide is formed when carbon burns in a limited amount of air or oxygen.

$$\text{carbon} + \text{oxygen (excess)} \xrightarrow{\Delta} \text{carbon dioxide}$$

$$C(s) + O_2(g) \xrightarrow{\Delta} CO_2(g)$$

$$\text{carbon} + \text{oxygen (limited)} \xrightarrow{\Delta} \text{carbon monoxide}$$

$$2C(s) + O_2(g) \xrightarrow{\Delta} 2CO(g)$$

(ii) Carbon can be used as a reducing agent in the extraction of some metals. When strongly heated, carbon can reduce the oxides of zinc and other metals such as CuO, PbO and Fe_2O_3 to their respective metals.

$$\text{carbon} + \text{zinc oxide} \xrightarrow{\Delta} \text{zinc} + \text{carbon monoxide}$$

$$C(s) + ZnO(s) \xrightarrow{\Delta} Zn(s) + CO(g)$$

Review Questions

(1) Diamond is very hard whereas graphite is soft. Why?
(2) Graphite is a good conductor of electricity, but diamond is not. Why?
(3) Explain why fullerene is an allotrope of carbon.
(4) Discuss the differences between graphite and graphene in their structures.

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Chemistry in Society

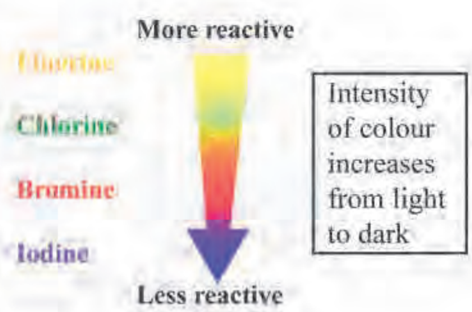
- Carbon is a major component of giant molecules called macromolecules which include proteins, lipids, nucleic acids and carbohydrates.
- There are different uses depending on the allotropes of carbon in everyday life.
- Diamond is the hardest substance known. Diamonds are used as glass cutters and drill points. Diamond is used for jewellery because of its brilliant shine.
- Graphite is used in lead pencil as it is soft. Powdered graphite is used as dry lubricant for machine parts which operate at high temperature where oil cannot be used because graphite is non-volatile. It is used in making electrode in the cells. Graphite crucibles are used as containers for melting metals at high temperature.
- Fullerene is used in artificial photosynthesis, cosmetics, surface coating of medical devices and drug delivery system.
- Graphene is widely used for solar cells, light-emitting diodes (LED), touch panels and smart windows or phones.
- Carbon (very small amount) is used to make some types of steel. Charcoal is used as a fuel for cooking. Activated charcoal is used as adsorbent in industry for bleaching (removal of colour), deodourization (removal of smell) of substances and in water purification. Coal is used as a fuel and also used to produce coke and coal tar. Coke is used as a fuel in metal industry and as reducing agent in the extraction of metals (lead, iron and zinc, etc.). Carbon black is used for making printing ink, black shoes polish and as filler in vehicle tyres and other rubber products.

Key Term

- **Allotrope** refers to two or more forms of an element that occur in the same physical state but different in properties.

5.3 HALOGENS

Halogens (F_2, Cl_2, Br_2, I_2) are diatomic molecules. Fluorine and chlorine are gases, bromine is a liquid and iodine is a solid at room temperature. They are electronegative elements. Since essential electronic structure of halogen is $ns^2\ np^5$, they are very reactive. Among them fluorine is the most reactive. Thus, none of the halogens can be found in nature in their elemental forms. They are found as salts of the halides.



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(a) Preparation of Halogens

In the laboratory, halogen is prepared by heating sodium or potassium halide with manganese(IV) oxide and concentrated sulphuric acid.

Activity (2): Preparation of chlorine

Chlorine is prepared from the mixture of sodium chloride, manganese(IV) oxide and concentrated sulphuric acid on heating. It is collected by the upward displacement of air because it is slightly soluble in water and heavier than air. Due to its greenish yellow colour, it can be seen easily when the gas jar is full of chlorine gas. If required dry, the gas is passed into concentrated sulphuric acid.

sodium chloride + maganese(IV) oxide + sulphuric acid(conc.) $\xrightarrow{\Delta}$ sodium hydrogen sulphate + manganese(II) sulphate + chlorine + water

$$2NaCl(s) + MnO_2(s) + 3H_2SO_4(l) \xrightarrow{\Delta} 2NaHSO_4(aq) + MnSO_4(aq) + Cl_2(g) + 2H_2O(l)$$

(b) Properties of Halogens

Physical properties

Fluorine is a very pale yellow gas. It is the lightest halogen and exists as a highly toxic gas. As it is the most electronegative element, it is extremely reactive. It reacts with almost all other elements, except helium and neon.

Chlorine is a pale green gas with a choking, unpleasant smell. Chlorine is very poisonous if inhaled even in small quantities. One part of chlorine in 50,000 parts of air may be harmful. Chlorine is about 2½ times as dense as air.

Bromine is a heavy, reddish brown, volatile liquid. It has a choking, irritating smell. Bromine means 'a stench'. Liquid bromine causes burns on the flesh, which heal with difficulty. Bromine is slightly soluble in water, forming a yellowish red solution containing about 3 percent of bromine at ordinary temperature.

Iodine is a purple-black shiny solid and irritating smell. Iodine sublimes rapidly when heated, forming a violet vapour from which the black solid can again be obtained by cooling. Iodine is almost insoluble in water but readily dissolves in aqueous potassium iodide. This is due to the formation of a compound of potassium iodide and iodine, which is very soluble. This solution is brown. Iodine also dissolves in ethanol and ether, forming brown solutions, and in carbon disulphide and carbon tetrachloride, forming violet solutions.

Chemical properties

Reactivity of halogens occurs in the order of $F_2 > Cl_2 > Br_2 > I_2$.

(i) Affinity for hydrogen

Halogens react with most compounds containing hydrogen.

hydrogen sulphide + chlorine $\longrightarrow$ hydrogen chloride + sulphur

$$H_2S(g) + Cl_2(g) \longrightarrow 2HCl(g) + S(s)$$

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When a tube containing equal volumes of chlorine and hydrogen is exposed to sunlight, it explodes.

$$\text{hydrogen} + \text{chlorine} \xrightarrow{\text{sunlight}} \text{hydrogen chloride}$$
$$H_2(g) + Cl_2(g) \xrightarrow{\text{sunlight}} 2HCl(g)$$

Bromine combines with hydrogen, but not as readily as chlorine does. A mixture of bromine and hydrogen needs heat to make them combine. Also, the hydrogen bromide formed is not as stable as hydrogen chloride. Iodine has little affinity for hydrogen.

(ii) Action with metals

Chlorine reacts vigorously with metals to form metal chlorides.

$$\text{iron} + \text{chlorine} \longrightarrow \text{iron(III) chloride}$$
$$2Fe(s) + 3Cl_2(g) \longrightarrow 2FeCl_3(s)$$

When a very thin sheet of an alloy of copper and zinc, mainly copper, is dropped into a chlorine gas jar, it burns brightly with a green flame.

$$\text{copper} + \text{chlorine} \longrightarrow \text{copper(II) chloride}$$
$$Cu(s) + Cl_2(g) \longrightarrow CuCl_2(s)$$
$$\text{zinc} + \text{chlorine} \longrightarrow \text{zinc chloride}$$
$$Zn(s) + Cl_2(g) \longrightarrow ZnCl_2(s)$$

Bromine combines readily with most metals to form bromides. For example, copper, iron and sodium give the corresponding bromides.

Iodine is fairly active and will combine with many metals to form iodides, but it does so, much less readily than either chlorine or bromine.

(iii) Action with non-metals

Phosphorus burns spontaneously in chlorine.

$$\text{phosphorus} + \text{chlorine} \longrightarrow \text{phosphorus(III) chloride}$$
$$2P(s) + 3Cl_2(g) \longrightarrow 2PCl_3(l)$$
$$\text{phosphorus} + \text{chlorine} \longrightarrow \text{phosphorus(V) chloride}$$
$$2P(s) + 5Cl_2(g) \longrightarrow 2PCl_5(s)$$

Bromine explodes when mixed with yellow phosphorus. Phosphorus(III) bromide is made by gradually adding a solution of bromine in carbon tetrachloride to red phosphorus.

$$\text{red phosphorus} + \text{bromine} \longrightarrow \text{phosphorus(III) bromide}$$
$$2P(s) + 3Br_2(g) \longrightarrow 2PBr_3(l)$$

(iv) Bleaching action

Among halogens, chlorine and bromine have bleaching power, however, iodine does not. Chlorine reacts with water to form HOCl and HCl.

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chlorine + water ⟶ hypochlorous acid + hydrochloric acid

$$Cl_2(g) + H_2O(l) \longrightarrow HOCl(aq) + HCl(aq)$$

The HOCl slowly decomposes to nascent oxygen.

hypochlorous acid ⟶ hydrochloric acid + nascent oxygen

$$2HOCl(aq) \longrightarrow 2HCl(aq) + 2[O]$$

This process is speeded up by light. The nascent oxygen bleaches dyes by oxidising them. Chlorine will bleach moist litmus papers. This is used as a test for chlorine. Bromine is also used as a bleaching agent but it is not as effective as chlorine. Bromine also bleaches moist litmus papers.

(v) Oxidising properties

Halogens are oxidising agents.
Chlorine oxidises iron(II) chloride into iron(III) chloride.

iron(II) chloride + chlorine ⟶ iron(III) chloride

$$2FeCl_2(aq) + Cl_2(g) \longrightarrow 2FeCl_3(aq)$$

Bromine is also an oxidising agent, but it is not as strong an oxidising agent as chlorine. Bromine will also give majority of oxidation reactions given by chlorine.

Iodine is a mild oxidising agent, but it will not perform many of the ordinary oxidising actions attributed to chlorine and bromine. However, iodine oxidises hydrogen sulphide to form hydrogen iodide and liberate sulphur.

hydrogen sulphide + iodine ⟶ hydrogen iodide + sulphur

$$H_2S(g) + I_2(g) \longrightarrow 2HI(g) + S(s)$$

(vi) Reaction with alkalis

Halogens also react with alkalis.

Chlorine reacts on solutions of alkalis in the same way as bromine. The actions of bromine and iodine on alkalis are similar to that of chlorine.

potassium hydroxide (cold dil.) + chlorine ⟶ potassium chloride + potassium hypochlorite + water

$$2KOH(aq) + Cl_2(g) \longrightarrow KCl(aq) + KOCl(aq) + H_2O(l)$$

potassium hydroxide (hot conc.) + chlorine ⟶ potassium chloride + potassium chlorate + water

$$6KOH(aq) + 3Cl_2(g) \longrightarrow 5KCl(aq) + KClO_3(aq) + 3H_2O(l)$$

(vii) Displacement properties

When chlorine is passed into a solution of sodium bromide or sodium iodide, the respective halogen is displaced by more reactive chlorine.

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chlorine + sodium bromide $\longrightarrow$ sodium chloride + bromine

$$Cl_2(g) + 2NaBr(aq) \longrightarrow 2NaCl(aq) + Br_2(aq)$$

When bromine is passed into a solution of sodium iodide, iodine is liberated. However, iodine cannot displace chlorine and bromine from their salt solutions.

(c) Halides

A halide ion is a halogen atom with a negative charge. The halide anions are fluoride (F^-), chloride (Cl^-), bromide (Br^-) and iodide (I^-).

Test for halides

The presence of chloride, bromide or iodide ions can be tested by adding silver nitrate solution (Figure 5.5). Samples are typically acidified with dilute nitric acid to remove interfering ions, e.g. carbonate ions. Different colours of silver halides precipitates are observed.

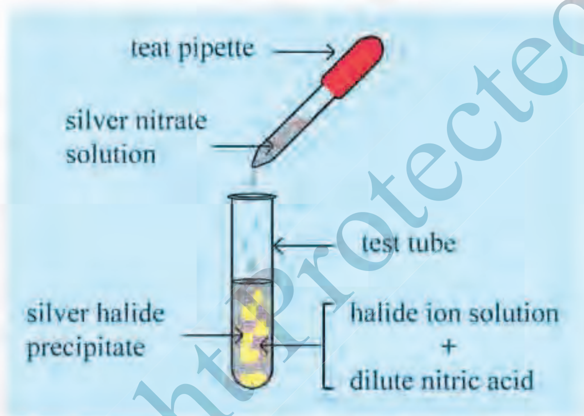

Figure 5.5 Silver Nitrate Test for Halide Ions

Chemistry in Society

- Fluorine is used in the form of fluorides in drinking water and toothpaste. It reduces tooth decay by hardening the enamel on teeth.
- Chlorine is used to make PVC (Polyvinyl chloride) plastic as well as household bleaches. It is also used to kill bacteria and viruses in drinking water.
- Bromine is used to make disinfectants, medicines and fire retardants.
- Iodine is used in medicines (e.g., to treat cases of goiter) and disinfectants (due to its antiseptic properties. e.g. 'Tincture of iodine') and also as a photographic chemical.
- Halides are used in the solder paste. It is widely used in metal halide lamps that are high-intensity discharge lamps.

Review Questions

(1) What does the term 'halogen' mean?
(2) Why are halogens highly reactive?
(3) What compound is formed when chlorine is passed over heated iron? What property does chlorine show in this reaction?
(4) Bromine reacts with sodium iodide. What property would you expect this bromine to have?
(5) What property does iodine show in the reaction with hydrogen sulphide?

Key Terms

- Siliver nitrate test is used to detect the presence of halide ions.

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EXERCISES

1. Write TRUE or FALSE for each of the following statements. If FALSE, correct it.
   (a) Oxygen will burn but it cannot support combustion.
   (b) Carbon cannot exhibit allotropy or polymorphism.
   (c) Fullerene is the synthetic allotrope of carbon.
   (d) Halogens are usually found as metal halides.
   (e) Bromine can displace the chlorine from metal chloride.
   (f) Sulphur dioxide is used as fire extinguishers.
   (g) All halogens have similar properties but not identical.
2. Fill in the blanks with a suitable word or words.
   (a) When sulphur is burnt in oxygen, _____ is obtained.
   (b) Oxygen is used as a / an _____ flame in the cutting and welding of steel.
   (c) An acidic oxide is a _____ oxide.
   (d) Diamond, graphite, fullerene and graphene are the same _____.
   (e) An allotrope of carbon, _____ can be used in surface coating of medical devices.
   (f) A disease, goiter, is caused due to deficiency of _____.
   (g) In halogens, _____ is a liquid at room temperature.
3. Write equations for the following reactions that show carbon dioxide as an acidic oxide.
   (a) reaction with water (b) reaction with sodium hydroxide solution
4. Give two examples to show that carbon has reducing properties.
5. What happens when
   (a) chlorine is passed into a potassium iodide solution?
   (b) a very thin sheet of an alloy of copper and zinc is dropped into a chlorine gas jar?
   (c) chlorine is passed over heated iron?
   (d) bromine reacts with cold dilute sodium hydroxide solution?
   (e) iodine vapour is passed into hydrogen sulphide?
6. Write chemical equations in words and symbols for the following chemical reactions:
   (a) Oxidation reaction of chlorine
   (b) Displacement reaction of bromine
   (c) Affinity for hydrogen on iodine
   (d) Reaction of bromine with concentrated potassium hydroxide solution
   (e) Reaction of sodium iodide with silver nitrate solution
   (f) Reaction of iodine with dilute potassium hydroxide solution
7. Write chemical equations for the preparation of bromine and iodine in laboratory.
8. How can you test the presence of chloride, bromide or iodide in a solution?
9. Halogens are strong oxidising agents. Explain with chemical equations.
10. Why is chlorine added to swimming pool water?

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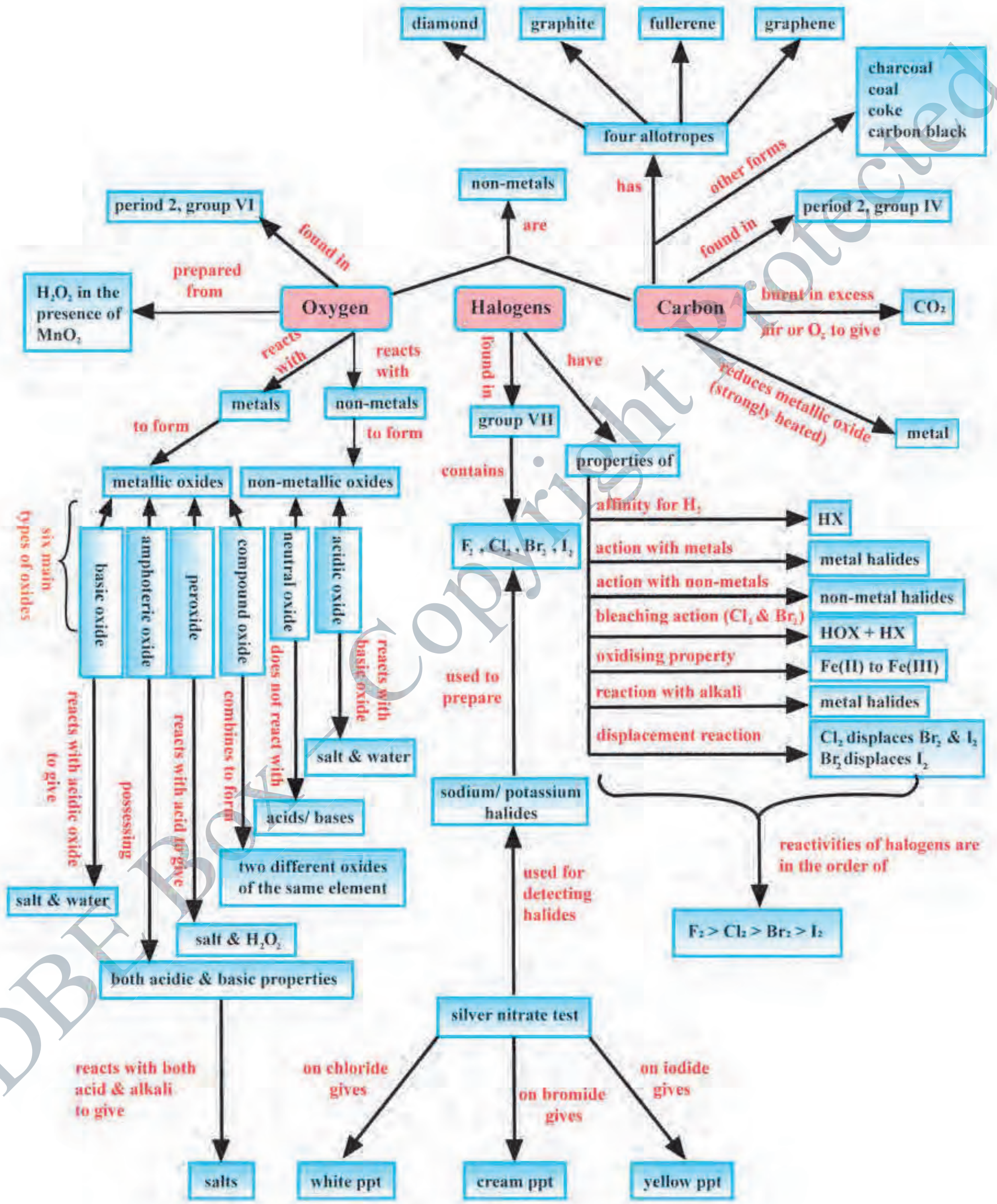
CHAPTER REVIEW
(Concept Map)



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CHAPTER 6

ACIDS, BASES AND SALTS

Acid-Base chemistry is important in a wide variety of everyday life. In our bodies, in our home and in our industrial society, acids, bases and salts play key roles.

In our bodies, proteins, enzymes, blood, genetic materials and other components of living matter contain both acids and bases.

The organs of human and animals also contain acids. You probably know how painful a bee sting or an ant bite can be. The pain is caused by an acid called methanoic (formic) acid. The pain we sometimes feel in our leg muscles during exercise is caused by lactic acid. Our stomach produces an acid (HCl) for food digestion.

Most important mineral acids such as sulphuric acid, hydrochloric acid and nitric acid are used in industries and laboratories. In our home, many cleaners contain acids or bases. For instance, a floor cleaner acid often contains hydrochloric acid and a glass cleaner base is ammonium hydroxide. Ordinary battery acid is sulphuric acid.

H_2SO_4 (battery acid)

HCl

HNO_3

CH_3COOH (vinegar)

Antacid tablets

NH_4OH (glass cleaner)

Learning Outcomes

After completing this chapter, students will be able to:

- describe the physical and chemical properties of acids, bases, alkalis and salts and their uses in daily life;
- distinguish between bases and alkalis;
- relate the role of indicators and the pH scale;
- classify the salts based on acids used and describe the preparation of salts;
- distinguish between soluble salts and insoluble salts.

We often use salts in our home. We sprinkled sodium chloride on our food to bring out its taste. We may use bath salts to help us relax in the bath and some of the medicines we take are salts. Salts are used as a preservative in pickles and in curing meat and fish, in the manufacture of soap, keeping ice from melting and making chemicals like washing soda, baking soda, etc.

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Svante Arrhenius Theory (1887)

Acid is a substance which when dissolved in water produces hydrogen ions (H^+). In other words, an acid increases the number of H^+ ions in an aqueous solution.

Base is a substance which when dissolved in water produces hydroxide ions (OH^-). In other words, a base increases the concentration of OH^- ions in an aqueous solution.

6.1 ACIDS AND THEIR PROPERTIES

Many 'acids' are corrosive, meaning they destroy body tissue and clothing and many are also poisonous. Acids can be found in many foods we eat. Some organic acids are used in food preservative, food fermentation, salad, etc., such as ethanoic acid (acetic acid). Some organic acids are found in the food presented in Figure 6.1.

Yogurt (lactic acid) | Apples (malic acid) | Citrus fruits (citric acid) | Grapes (tartaric acid) | Gooseberries (Zee Phyu) (ascorbic acid) | Roselle plant (Chin Baung) (oxalic acid)

Figure 6.1 Occurrence of Some Organic Acids in Nature

(a) Acids

The word 'acid' comes from the Latin word *acidus*, which means sour. Acids can be classified as mineral acids (inorganic acids) and organic acids. Acids can be strong or weak. Strong acid cannot dissociate itself without water. When a strong acid dissolves in water, it completely dissociates to produce hydrogen ions, which are protons, (H^+), and a weak acid only partially dissociates in water. Mineral acids are strong acids and organic acids are weak acids. Some acids and their dissociation reactions in water are described in Table 6.1.

Table 6.1 Names and Formulae of Some Acids and Their Dissociation Reactions in Water

Name of acids	Chemical formula	Dissociation reaction in water	Strength of acids
hydrochloric acid	HCl	$HCl(aq) \longrightarrow H^+(aq) + Cl^-(aq)$	strong
sulphuric acid	H_2SO_4	$H_2SO_4(aq) \longrightarrow 2H^+(aq) + SO_4^{2-}(aq)$	strong
nitric acid	HNO_3	$HNO_3(aq) \longrightarrow H^+(aq) + NO_3^-(aq)$	strong
ethanoic acid	CH_3COOH	$CH_3COOH(aq) \rightleftharpoons CH_3COO^-(aq) + H^+(aq)$	weak

$\longrightarrow$ completely dissociates in water $\rightleftharpoons$ partially dissociates in water

(**Caution**: Always add strong acid slowly to water. This is because the acid becomes very hot and splashing may happen.)

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Properties

An acid is a compound which becomes a proton (H^+) donor when dissolved in water. The properties and reactions of an acid are due to these hydrogen ions.

Physical properties

(i) Acids are hazardous, irritant and corrosive.

(ii) Acids have a sour taste.
(DON'T TASTE, DON'T TOUCH.)

(iii) Acids dissolve in water to form solutions which conduct electricity.

(iv) Acid solutions have pH values less than 7.

(v) Acids have the ability to change the colour of indicators and turn blue litmus paper (an indicator) red.

Chemical properties

(i) Acid reacts with metals to form a salt and hydrogen.

$$\boxed{\text{metal} + \text{acid (dil.)} \longrightarrow \text{salt} + \text{hydrogen}}$$

$$\text{calcium} + \text{hydrochloric acid (dil.)} \longrightarrow \text{calcium chloride} + \text{hydrogen}$$
$$Ca(s) + 2HCl(aq) \longrightarrow CaCl_2(aq) + H_2(g)$$

(ii) Acid reacts with metal oxides and hydroxides to form a salt and water only.

$$\boxed{\text{metal oxide} + \text{acid (dil.)} \longrightarrow \text{salt} + \text{water}}$$

$$\text{calcium oxide} + \text{hydrochloric acid (dil.)} \longrightarrow \text{calcium chloride} + \text{water}$$
$$CaO(s) + 2HCl(aq) \longrightarrow CaCl_2(aq) + H_2O(l)$$

$$\boxed{\text{metal hydroxide} + \text{acid (dil.)} \longrightarrow \text{salt} + \text{water}}$$

$$\text{calcium hydroxide} + \text{sulphuric acid (dil.)} \longrightarrow \text{calcium sulphate} + \text{water}$$
$$Ca(OH)_2(s) + H_2SO_4(aq) \longrightarrow CaSO_4(s) + 2H_2O(l)$$

(iii) Acid reacts with carbonate to form a salt, carbon dioxide and water.

$$\boxed{\text{metal carbonate} + \text{acid (dil.)} \longrightarrow \text{salt} + \text{carbon dioxide} + \text{water}}$$

$$\text{calcium carbonate} + \text{hydrochloric acid (dil.)} \longrightarrow \text{calcium chloride} + \text{carbon dioxide} + \text{water}$$
$$CaCO_3(s) + 2HCl(aq) \longrightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$$

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Chemistry in Daily Life

Some examples of most common uses of acids in daily life are listed in the following Table:

Acids	Formula	Uses
sulphuric acid	H_2SO_4	extraction of some metals such as copper, manufacture of fertilisers, detergents, paints, rubber, paper and pulp industry, car batteries and rust removal
hydrochloric acid	HCl	to help swimming pools be free of algae, to make aqua regia for dissolving gold and platinum
nitric acid	HNO_3	making fertilisers and explosives, to make aqua regia (a mixture of one part of the concentrated nitric acid and three parts of the concentrated hydrochloric acid) for dissolving gold and platinum
phosphoric acid	H_3PO_4	making fertilisers and rust inhibitor
carbonic acid	H_2CO_3	in fizzy drinks
citric acid	$C_6H_8O_7$	in fruit juices, in the preparation of effervescent salts, as a food preservative
ethanoic acid	CH_3COOH	in vinegar, used in salad dressings

Review Questions

(1) After rubbing an old copper coin with lemon juice, what visible change happens to the coin? Why?
(2) How can you detect whether a solution is acidic or not? **(Not to taste)**
(3) Ant bite is painful. Why is it so?
(4) Why can you treat bee stings with baking powder?
(5) In a laboratory, solution **A** is prepared by dissolving 10 mL of hydrochloric acid in 100 mL of water and solution **B** is prepared by dissolving 1 mL of hydrochloric acid in 100 mL of water. Which one is more concentrated? Which one is strong or weak or not?

Key Terms

- An **acid** is a compound that dissolves in water to produce hydrogen ions, H^+.
- A **dissociation reaction** is a chemical reaction in which a compound breaks apart into two or more parts.
- **Strong acid** is an acid that completely dissociates in water and gives H^+ ions. All strong acid molecules become ions in the water.

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- **Weak acid** is an acid that partially dissociates in aqueous solution and gives H^+ ions. Most of these molecules remain unchanged in the water.
- An acid solution which contains the pure acid or predominantly large proportion of the acid is called **concentrated acid**.
- An acid solution which contains a relatively small amount of the acid is called **dilute acid**.

6.2 BASES, ALKALIS AND THEIR PROPERTIES

Bases and alkalis are found in many cleaning agents such as soap and many household detergents. When wood ashes are burnt, the product is alkaline. The word **alkali** comes from the Arabic 'al-qili' which means burnt ashes. It is used traditionally by gardeners as a good source of potash.

(a) Bases

A base is usually a metallic oxide or hydroxide and will react with an acid to form a salt and water only. For example,

$$\text{copper(II) oxide} + \text{sulphuric acid (dil.)} \longrightarrow \text{copper(II) sulphate} + \text{water}$$
$$CuO(s) + H_2SO_4(aq) \longrightarrow CuSO_4(aq) + H_2O(l)$$

$$\text{sodium hydroxide} + \text{hydrochloric acid (dil.)} \longrightarrow \text{sodium chloride} + \text{water}$$
$$NaOH(aq) + HCl(aq) \longrightarrow NaCl(aq) + H_2O(l)$$

(b) Alkalis

An alkali is a base that is soluble in water. An example of a soluble base is sodium oxide.

$$\text{sodium oxide} + \text{water} \longrightarrow \text{sodium hydroxide}$$
$$Na_2O(s) + H_2O(l) \longrightarrow 2NaOH(aq)$$

Alkalis can be strong or weak. Strong alkalis dissolve in water to produce OH^- ions in solution. Sodium hydroxide and potassium hydroxide are examples of strong alkalis. Ammonium hydroxide is the most common example of a weak alkali.

$$\text{sodium hydroxide} \longrightarrow \text{sodium ion} + \text{hydroxide ion}$$
$$NaOH(aq) \longrightarrow Na^+(aq) + OH^-(aq)$$

$$\text{ammonium hydroxide} \rightleftharpoons \text{ammonium ion} + \text{hydroxide ion}$$
$$NH_4OH(aq) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$$

Most bases are insoluble in water. MgO, CuO, Fe_2O_3, etc. are insoluble bases. They do not react with water and also not dissolve in water. Thus, it is a base and not an alkali. Some common alkalis and bases are described in Table 6.2.

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Table 6.2 Some Common Alkalis and Bases

Name	Formula	Alkalis (soluble bases)	Insoluble bases
sodium oxide	Na_2O	alkali	-
sodium hydroxide	$NaOH$	alkali	-
potassium oxide	K_2O	alkali	-
potassium hydroxide	KOH	alkali	-
calcium oxide	CaO	alkali	-
calcium hydroxide	$Ca(OH)_2$	alkali	-
copper(II) oxide	CuO	-	base
magnesium oxide	MgO	-	base
iron(III) oxide	Fe_2O_3	-	base

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Properties

Physical properties

(i) Strong bases are hazardous to handle.
(ii) Bases have a bitter taste and soapy feel. **(DON'T TASTE)**
(iii) Bases cause a colour change in indicators. Litmus changes from red to blue in a basic solution.
(iv) Alkalis have pH values greater than 7.

Chemical properties

(i) Bases react with acids to neutralise each other and form a salt and water. For example,

magnesium oxide + sulphuric acid ⟶ magnesium sulphate + water

$$MgO(s) + H_2SO_4(aq) \longrightarrow MgSO_4(aq) + H_2O(l)$$

sodium hydroxide + sulphuric acid ⟶ sodium sulphate + water

$$2NaOH(aq) + H_2SO_4(aq) \longrightarrow Na_2SO_4(aq) + 2H_2O(l)$$

(ii) When alkalis are gently warmed with ammonium salt it gives off ammonia gas.

sodium hydroxide + ammonium chloride $\xrightarrow{\Delta}$ sodium chloride + water + ammonia

$$NaOH(aq) + NH_4Cl(s) \xrightarrow{\Delta} NaCl(aq) + H_2O(l) + NH_3(g)$$

(iii) Alkalis react with fatty acids to form soaps.

Chemistry in Daily Life

Some common bases and alkalis and their uses are described in the following Table:

Bases and alkalis	Formula	Uses
sodium hydroxide	$NaOH$	making soap, paper, baking soda, oven cleaners
calcium hydroxide (slaked lime)	$Ca(OH)_2$	treating acidic soil (liming), making cement, limewater, mortar, plaster
calcium oxide (quicklime)	CaO	making cement
magnesium oxide	MgO	in antacids (gastric medicine), in toothpaste
ammonia	NH_3	in many household cleaners and production of fertilisers

Review Questions

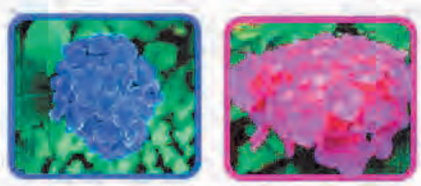
(1) Oven cleaner can remove the dirt and grease from oven. What is the active ingredient in the cleaner? What is the function of that ingredient?

(2) Toothpaste contains aluminium hydroxide which removes plague. What does this tell you about the nature of plaque and bacteria on your teeth?

Key Terms

- A **base or** an **alkali** is a chemical compound that combines with an acid to form a salt and water. An **alkali** is a base which is soluble in water producing OH^- ions. All **alkalis** are bases but all bases are not alkalis.
- A **strong base** is a base that completely dissociates in water producing OH^- ions. All base molecules become ions in the water.
- A **weak base** is a base that partially dissociates in water producing OH^- ions. Most of the base molecules remain unchanged in the water.

6.3 INDICATORS AND THE pH SCALE



Many brightly coloured flowers, vegetables and berries make good indicators. For example, the coloured juice extracted from red cabbage is pink in acids and green in alkalis. Hydrangea flowers are interesting natural indicators. They are blue when grown in acidic soil and pink or red when grown in alkaline soil.

(a) Indicators

Indicators are dyes, or a mixture of dyes, which change colour when they are added to acids or alkalis. Some indicators can be used to determine pH because of their colour changes somewhere along the pH scale (Figure 6.2). Litmus is red in acidic solution, purple in neutral and blue in alkaline solution.

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(b) The pH Scale

A measure of the acidity or alkalinity of the solution is known as pH. The pH value can be measured by pH meter (Figure 6.3). It is a much more reliable and accurate method of measuring pH than the universal indicator paper.

Substances in the body have different pH values. Acidic conditions in the stomach (pH~1.5) are needed for good digestion. Usually the body maintains the pH of blood close to 7.4. The pH scale demonstrates the strength of an acid or alkali (Figure 6.2). Solutions and their pH values are described in Table 6.3.

Figure 6.2 The pH Scale

Table 6.3 Solutions and Their pH Values

Solutions	pH value (0 to 14)
acidic	below 7
basic	above 7
neutral	equal to 7

(a) (b)

Figure 6.3 (a) The pH Meter
(b) Universal Indicator Paper

Chemistry in Daily Life

- The pH is important for the correct functioning of the body, for food and water and for the growth of plants.
- Many plants do not grow properly in highly acidic or highly alkaline soil. Highly acidic soil is treated by spreading quicklime (CaO), slaked lime ($Ca(OH)_2$) or calcium carbonate ($CaCO_3$) to lower its acidity.
- Highly alkaline soil is treated by adding gypsum ($CaSO_4.2H_2O$) to lower its alkalinity.

Review Questions

(1) Which of the solutions having the following pH, are acidic or alkaline or neutral?
(a) pH 6 (b) pH 3 (c) pH 7 (d) pH 8
(2) The pH of pancreatic juice is 7.9. Is pancreatic juice acidic or basic?
(3) How do we detect whether a soil is acidic or basic?
(4) Name a common household substance with a pH (a) greater than 7 (b) less than 7 (c) almost 7.

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Key Terms

- An **indicator** is a substance that has different colours in acidic and alkaline solutions.
- A measure of the acidity or alkalinity of a solution is known as its **pH**. Solutions with pH < 7 are acidic and those with pH > 7 are alkaline. The solutions of pH 7 are neutral. The pH of pure water is 7. The pH of a solution can be measured by using the pH meter.

6.4 SALTS

Many different types of salts can be found in nature. The sea water contains many salts such as sodium chloride, potassium chloride, magnesium sulphate, magnesium chloride and magnesium bromide.

The Earth's crust is made up of minerals containing various types of salts such as calcium fluoride (fluorite), magnesium sulphate (Epsom salt), lead(II) sulphide (galena) and calcium carbonate (limestone), etc.

A salt is produced when an acid reacts with a base. The salt consists of two parts. One part comes from the base, the other from the acid. An example is sodium chloride, NaCl, produced from sodium hydroxide and hydrochloric acid.

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

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

The sodium ion (Na^+) part of the salt comes from the base and the chloride ion (Cl^-) comes from the acid. When an acid reacts with a base, a salt and water are formed. This reaction is known as **neutralisation**. It involves the combination of H^+ ion produced from acid and OH^- ion produced from base to form water.

$$\text{hydrogen ion} + \text{hydroxide ion} \longrightarrow \text{water}$$

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

Neutralisation reaction occurs in our stomach. The acid (HCl) produced by our stomach, is so strong that it is neutralised with a base produced by cells. Salts are also produced when an acid reacts with a metal or a metal carbonate.

(a) Classification of Salts

The salts can be classified based on acids used. Some examples of salts (chloride, sulphate, nitrate, sulphite and carbonate salts) formed from different acids are shown in Table 6.4.

Some salts are soluble and some are insoluble depending on the types of metals. The examples of soluble and insoluble salts are given in Table 6.5.

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Table 6.4 Some Salts Formed from Different Acids

Acids		Salts	
hydrochloric acid	HCl	**chloride salts** sodium chloride zinc chloride magnesium chloride	 NaCl $ZnCl_2$ $MgCl_2$
sulphuric acid	H_2SO_4	**sulphate salts** sodium sulphate copper(II) sulphate	 Na_2SO_4 $CuSO_4$
nitric acid	HNO_3	**nitrate salts** sodium nitrate potassium nitrate ammonium nitrate copper(II) nitrate	 $NaNO_3$ KNO_3 NH_4NO_3 $Cu(NO_3)_2$
sulphurous acid	H_2SO_3	**sulphite salts** sodium sulphite	 Na_2SO_3
carbonic acid	H_2CO_3	**carbonate salts** sodium carbonate calcium carbonate	 Na_2CO_3 $CaCO_3$
ethanoic acid	CH_3COOH	**ethanoate salt** sodium ethanoate	 CH_3COONa

Table 6.5 Soluble and Insoluble Salts

Salts	Soluble salts	Insoluble salts
nitrates	all nitrates	none
chlorides	all chlorides (except silver, mercury(I), lead(II))	silver, mercury(I), lead(II)
sulphates	all sulphates (except barium, lead(II), calcium)	barium, lead(II), calcium
carbonates	sodium, potassium, ammonium	all carbonates except those of sodium, potassium and ammonium

(b) Preparation of Salts

The preparation method depends on whether the salt is soluble in water or not. Soluble salts are usually prepared by crystallisation method. Insoluble salts are usually prepared by precipitation method.

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Soluble salts

(i) Soluble salts may be prepared by using moderately dilute acids and metals. The salt formed can then be separated by crystallisation. For example,

zinc + sulphuric acid (dil.) ⟶ zinc sulphate + hydrogen

$$Zn(s) + H_2SO_4(aq) \longrightarrow ZnSO_4(aq) + H_2(g)$$

(ii) Salts of sodium, potassium and ammonium can be prepared from caustic soda solution (NaOH), caustic potash solution (KOH) and ammonia solution (NH_4OH), respectively, by the neutralisation using the appropriate acid.

sodium hydroxide + hydrochloric acid (dil.) ⟶ sodium chloride + water

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

potassium hydroxide + sulphuric acid (dil.) ⟶ potassium sulphate + water

$$2KOH(aq) + H_2SO_4(aq) \longrightarrow K_2SO_4(aq) + 2H_2O(l)$$

ammonium hydroxide + nitric acid (dil.) ⟶ ammonium nitrate + water

$$NH_4OH(aq) + HNO_3(aq) \longrightarrow NH_4NO_3(aq) + H_2O(l)$$

(iii) Soluble salts can be prepared by using either the oxide or the hydroxide of the metal with the appropriate acid.

zinc oxide + sulphuric acid (dil.) ⟶ zinc sulphate + water

$$ZnO(s) + H_2SO_4(aq) \longrightarrow ZnSO_4(aq) + H_2O(l)$$

zinc hydroxide + sulphuric acid (dil.) ⟶ zinc sulphate + water

$$Zn(OH)_2(s) + H_2SO_4(aq) \longrightarrow ZnSO_4(aq) + 2H_2O(l)$$

Insoluble salts

Insoluble salts are prepared by precipitation. For example, an insoluble salt, barium sulphate, can be made by mixing solutions of barium chloride and potassium sulphate. A white precipitate of barium sulphate, $BaSO_4$, is formed.

barium chloride + potassium sulphate ⟶ barium sulphate + potassium chloride

$$BaCl_2(aq) + K_2SO_4(aq) \longrightarrow BaSO_4(s) + 2\,KCl(aq)$$

Review Questions

(1) How would you neutralise hydrochloric acid if you spill it on the floor of a laboratory?
(2) If the soil is too acidic, we add lime to the soil. Explain the purpose of this.

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(3) Farmers treat the alkaline soil by using gypsum ($CaSO_4 . 2H_2O$). Why?
(4) We take gastric medicine when we feel stomach pain. Explain the action of this medicine.

Chemistry in Society

Salts play an important role in our society. Some salts and their uses in society are described in the following Table:

Salts	Formula	Uses
sodium chloride	$NaCl$	food additive
sodium sulphate sodium nitrite sodium citrate	Na_2SO_4 $NaNO_2$ $Na_3C_6H_5O_7$	food preservatives
ammonium sulphate ammonium nitrate ammonium phosphate	$(NH_4)_2SO_4$ NH_4NO_3 $(NH_4)_3PO_4$	fertilisers
potassium chloride	KCl	fertiliser
magnesium sulphate magnesium hydroxide	$MgSO_4.7H_2O$ $Mg(OH)_2$	medical uses (Epsom salt) medical uses (milk of magnesia, MOM)
calcium sulphate	$CaSO_4$	medical uses (Plaster of Paris , POP)

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Key Terms

- A **salt** is a substance produced from the reaction between an acid and a base or a metal. Based on the acids, salts can be classified as chlorides, sulphates, nitrates, sulphites and carbonates, etc. Soluble salts are usually prepared from the reactions between metals and dilute acids followed by crystallisation. Insoluble salts are usually prepared by precipitation method.
- **Neutralisation** is the reaction between an acid and a base to form a salt and water only.

EXERCISES

1. Think carefully about the following statements. Are they TRUE or FALSE? If FALSE, correct it.
   (a) In general, all acid solutions contain hydrogen ions, H^+.
   (b) Copper(II) hydroxide is an alkali.
   (c) The smaller the pH value, the more acidic a solution is.
   (d) Strong acids and alkalis are harmful and corrosive.
   (e) Litmus paper can measure the range of pH of a solution.

2. Select the correct word or words given in the brackets.
   (a) Complete the following equation:
   $2KOH + H_2SO_4 \longrightarrow ?$ ($KSO_4 + H_2O$; $K_2SO_4 + H_2O$; $KSO_4 + 2H_2O$; $K_2SO_4 + 2H_2O$)
   (b) Which of the following compounds can form an aqueous solution of pH >7? (carbon dioxide; hydrogen chloride; sodium chloride; sodium hydroxide)
   (c) Which of the following gases reacts with sulphuric acid to form a fertiliser? (ammonia; carbon dioxide; hydrogen; nitrogen)
   (d) A sample of pond water has a pH value of 11.
   This means that the water is (weakly acidic; neutral; weakly alkaline; strongly alkaline).
   (e) Which of the following substances could be used in excess to change the pH of soil from 5 to 7? (sodium chloride; calcium oxide; hydrochloric acid; sulphuric acid).

3. Fill in the blanks with a suitable word or phrase or numerical value with unit as necessary.
   (a) The combination of H^+ and OH^- ions to form water is called _____.
   (b) The pH of alkali solution is greater than _____.
   (c) Solutions having pH values below 4.5, turn blue litmus paper _____.
   (d) A measure of the acidity or alkalinity of a solution is known as _____.
   (e) The salt can be classified as soluble and insoluble salts depending on the types of ______.
   (f) Sodium citrate is a soluble salt. It is used as a food _____.

4. Complete the following sentences by using the words given below:
   (a) base, dissolves, hydrogen, ions, proton
   When an acid _____ in water, hydrogen _____are formed. A _____ ion is a proton. An acid is a_____ donor. It gives its proton to a _____.
   (b) hydroxides, hydrogen, dissolves, salt, oxides, water
   An acid is a compound that _____in water to produce _____ions. Acids react with metals to form _____and hydrogen. When acids react with metal _____ or _____, a salt and _____are formed.
   (c) acids, ammonia, hydroxide, salt, soluble
   Alkalis are water _____ bases. Examples of alkalis are_____ and sodium _____. Alkalis react with_____ to form a _____and water.
   (d) universal, alkaline, neutral, high, scale, seven, acidic
   The pH_____ shows how acidic or _____a solution is. Strongly_____ solutions have a low pH, strongly alkaline solution have a _____ pH. A solution that is neither acidic nor an alkaline is called a_____ solution. It has a pH of _____. The pH of a solution can be measured using _____indicator or a pH meter.

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