NGAIZA EDUCATION HUB (NEH)

Mastering Series

MASTERING CHEMISTRY CALCULATIONS FOR ORDINARY LEVEL

Author: NGAIZA, Lusima

Ngaiza education hub organisation
Mastering series

ISBN 978-9987-9677-2-8

Published by:

NEH organisation

P.O. BOX, 60791

Dar-es-salaam

+255748154159

PREFACE

Lack of books which give detailed explanation and enough number of worked examples to sharpen students' understanding on calculations involving chemistry subject for ordinary level of Tanzania secondary education is a big problem in the country. It is the long time problem to both teachers and students.

This chemistry book is a result of being aware and so much concerned to the problem. It is the result of having strong desire to find appropriate solution to the problem. My hope is that; this book will be part of the solution for the problem. This is sweet news to hear because the book discusses intensively all important concepts behind calculations accompanied with many worked examples which are carefully tactically and technically structured to enhance students' ability to tackle any calculation problem in ordinary level chemistry!

Organisation

The book is organised into ten chapters which discuss several matters in accordance to the ordinary level chemistry syllabus. Chapters are creatively ordered to blend together easy and deep understanding of matters.

At the end of each chapter there is an exercise which is termed as **practice exercise**. Both students and teachers will find these exercises very useful for assessing students' understanding of matters discussed in each respective chapter.

Near the end of the book there are **examination questions** which are very useful to students who wish to examine themselves on their understanding capacity of all matters discussed in the book. The questions cover knowledge of all topics discussed in the book; therefore, students are advised to attempt answering them after completing the study of all chapters. Also teachers will find the questions useful in examining students' understanding.

Nearer the end of the book, there are **answers and solutions for practice exercises and examination questions** for providing a guide to students on answering the questions. Students are recommended to check these solutions after an attempt of answering questions through writing down and not just through mind imagination of answers.

I hope both teachers and students will find this book helpful and interesting.

Acknowledgements

I am indebted to all students and teachers who were used to insist me about the need of preparing this book. Their belief on me provided good motivation in accomplishing difficult work of preparing this book.

I also wish to thank Ms. Monica Lema for typing the manuscript; Mr. Dickson Kihwelo for doing editorial work and typesetting of the manuscript and Ms. Salha Salum for doing very supportive editorial work.

Also, I thank Ms. Rahma Abdan for initiating the writing process, Mr. Canisius Joseph and Mr. Isakwisa Tende for their constructive ideas. Special thanks to Mrs. Gloria Lusima for all things she has done which was of great help for early accomplishment of this work.

Ngaiza Lusima

CONTENTS

Chapter one

OXIDATION STATE

INTRODUCTION

Oxidation state (also termed as **oxidation number**) *is the number of electrons gained or lost by an atom in combining with another atom by assumption that the bond formed between the atoms is ionic*.

- Oxidation number (**O.N**) gives the relative electron electronegativity of the two combining atoms; whereby the more electronegative atom has negative oxidation number while the more electropositive (less electronegative) has positive oxidation number.
- It is represented by positive or negative number (may be fraction or whole number) with negative or positive sign written before the number to differentiate from ionic charge whose sign is written after the number (which must be whole number).

RULES OF ASSIGNING OXIDATION NUMBERS

1. The oxidation number of any atom in neutral (free) element is zero. As an example, the oxidation number of carbon atom in elemental carbon is 0.
2. The total oxidation number of atoms in a neutral compound is zero. As an example, the summation of oxidation number of one O atom and two H atoms in H_2O is 0.
3. The oxidation number of an ion (charged atom) is equal to the charge of the ion. As an example, oxidation number of O atom in O^{2-} is -2.
4. The total oxidation number of atoms in a charged radical is equal to the amount of the charge of the radical. As an example, the summation of oxidation number of one C atom and three O atoms in the CO_3^{2-} is -2.
5. Some atoms of elements have fixed oxidation number in their compounds. For example:

- Atoms of all group I elements have oxidation number $+1$ in their compounds.
- Atoms of all group II elements have oxidation number of +2 in their compounds.
- Atoms of group III elements have oxidation number of +3 in their compounds.
- Fluorine atom has oxidation number of -1 in its compounds.

6. Oxidation number oxygen atom is -2 except in:

- Peroxides (e.g. H_2O_2, Na_2O_2) where its oxidation number is -1.
- F_2O where its oxidation number is +2
- Superoxides like KO_2 where is $-1/_2$.

7. Oxidation number of hydrogen atom is +1 except in ionic hydrides (e.g. NaH) where its oxidation number is -1.

WORKED EXAMPLES

Example 1

Find the oxidation state of sulphur (S) in the following chemical species:

(a) S_8
(b) S^{2-}
(c) SO_4^{2-}
(d) $S_4O_6^{2-}$

Solution

(a) S_8 is the **neutral molecule** of elemental sulphur and hence the oxidation state of sulphur will be 0.

(b) S^{2-} is the ionic form of sulphur and thus the oxidation of sulphur will be equal to the charge of the ion which is -2. Hence, the oxidation state of S is -2.

(c) SO_4^{2-} is the radical and thus the summation of oxidation states of atoms (one S and four O atoms) will be equal to the charge of the radical which is -2.
That is; $\mathrm{O.N\ of\ S + O.N\ of\ four\ O\ atoms} = -2$
But oxygen has fixed oxidation number of -2.
Then; $\mathrm{O.N\ of\ S} + (4 \times -2) = -2$
Or $\mathrm{O.N\ of\ S} + \ -8 = -2$
Or $\mathrm{O.N\ of\ S} = -2 + 8 = +6$
The oxidation state of S is $+6$.

(d) $S_4O_6^{2-}$ is the radical and thus the summation of oxidation states of atoms (four S and six O atoms) will be equal to the charge of the radical which is -2.
i.e. $\mathrm{O.N\ of\ four\ S\ atoms + O.N\ of\ six\ O\ atoms} = -2$
But oxygen has fixed oxidation number of -2.
Then; $4 \times \mathrm{O.N\ of\ S} + (6 \times -2) = -2$
Or $4 \times \mathrm{O.N\ of\ S} + \ -12 = -2$
Or $4 \times \mathrm{O.N\ of\ S} = -2 + 12 = +10$
Or $\mathrm{O.N\ of\ S} = \frac{+10}{4} = +2.5$
The oxidation state of S is $+2.5$.

Example 2

Identify oxidation number of each of the following atoms:

(a) $\mathrm{Cl\ in\ Cl_2}$
(b) $\mathrm{O\ in\ Cl_2O_5}$
(c) $\mathrm{Cl\ in\ NaOCl}$

Solution

(a) Cl_2 is the **neutral molecule** of elemental chlorine and hence the oxidation number of chlorine in Cl_2 is 0.

(b) Oxygen has fixed oxidation number of -2 and hence the oxidation number of O in Cl_2O_5 is -2.

(c) In this case:

- Na has fixed oxidation number of $+1$.
- O has fixed oxidation number of -2.
- Only Cl has no fixed oxidation number.
- Summation of oxidation numbers of all atoms in NaOCl is 0.

Let the oxidation number of Cl be x.

Then $(+1)+(-2)+x=0; x=+1$

Oxidation number of Cl is $+1$.

Example 3

Find oxidation state of each at atom in the following:

(a) KOH
(b) $NaClO_3$
(c) $Ca(NO_3)_2$

Solution

(a) In this case, all three atoms in the compound have fixed oxidation state and hence:

- Oxidation state of K is $+1$
- Oxidation state of O is -2
- Oxidation state of H is $+1$

(b) Here:

- Na has fixed oxidation state of $+1$.
- O has fixed oxidation state of -2.
- Only Cl has no fixed oxidation state.
- Summation of oxidation numbers of all atoms in $NaClO_3$ is 0.

Let the oxidation number of Cl be x.

Then $(+1)+(-2\times 3)+x=0; x=+5$

Oxidation state of Cl is $+5$.

(c) Here:

- Ca has fixed oxidation state of $+2$.
- O has fixed oxidation state of -2.
- Only N has no fixed oxidation state.
- Summation of oxidation numbers of all atoms in $Ca(NO_3)_2$ is 0.

Let the oxidation number of N be x.
Then $(+2) + (-2 \times 6) + 2x = 0; x = +5$
Oxidation state of N is $+5$.

Example 4

When a carbon atom is bonded to oxygen, it normally exhibits the valence of 4. The carbon atom may combine with three oxygen atoms to form carbonate radical. Based on this information, deduce the charge of the carbonate radical and hence its formula.

Solution

Since carbon is less electronegative than oxygen and it has valence of 4, it has positive oxidation state with value of 4. Thus the oxidation number of carbon in the carbonate radical is $+4$.

Since the oxygen atom has fixed oxidation state of -2; the three oxygen atoms is equivalent to oxidation state of $-2 \times 3 = -6$

So total oxidation state of one carbon atom and three oxygen atoms$= (+4) + (-6) = -2$.

Hence the carbonate radical has the charge of $2-$ with formula of CO_3^{2-}.

PRACTICE EXERCISE 1

Question 1

Find the oxidation state of Cl in each of the following:

(a) ClO^-
(b) $HClO_2$

Question 2

Deduce the oxidation state of the each of the following atoms:

(a) Fe in Fe^{3+}
(b) S in $MgSO_4$
(c) Mn in MnO_4^-
(d) H in $H_2C_2O_4$
(e) P in $Ca_3(PO_4)_2$

Question 3

Give oxidation state of each atom in the following compounds:

(a) $Ca(OH)_2$
(b) $Cu(NO_3)_2$
(c) NO

Question 4

Find oxidation state of sulphur atom in a compound which consists of two sodium atoms, two sulphur atoms and three oxygen atoms.

Question 5

Deduce the charge of the radical that is formed by combining one oxygen atom and one hydrogen atom.

Chapter two

AVERAGE ATOMIC MASS

INTRODUCTION

An atom *is the smallest individual particle of an element that can take part in a chemical reaction.*

- This means that the element is made from atoms or in imaginative language, we can say that; atoms are building blocks for the building called element.

These building blocks called atoms are made from smaller particles namely: **protons**, **neutrons** and **electrons** which are collectively known as **sub – atomic particles**.

- Among the three sub-atomic particles, by far proton and neutron are heavier than electron whose mass is negligible.
- These heavier particles which are found in very small core called nucleus are the one which contribute to the entire mass of the atom. Mathematically this can be written as follows:

Mass of an atom = total mass of protons + total mass of neutrons

Whereby:

Total mass of protons = mass of one proton × total number of protons (in the atom) and;

Total mass of neutrons = mass of one neutron × total number of neutrons (in the atom).

And mass of each proton and neutron is as follows:

Mass of one proton = 1amu

Mass of one neutron = 1amu

And of course mass of one electron = 0amu

Where, **amu** (or simply **u**) is the abbreviation for **a**tomic **m**ass **u**nit which is equivalent to 1.66×10^{-24}g or 1.66×10^{-27}kg.

That is 1amu = 1.66×10^{-24}g or 1amu = 1.66×10^{-27}kg.

Thus like proton and neutron, mass of an atom is commonly given in amu.

Example 1

X is the hypothetical atom whose number of protons and neutrons are 13 and 14 respectively. Calculate the approximated mass of one atom of **X**.

Solution

Using;

Mass of an atom = total mass of protons + total mass of neutrons

But:

Total mass of protons = total number of protons × mass of one proton

$$= 13 \times 1 \text{ amu} = 13 \text{ amu}$$

And total mass of neutrons = total number of neutrons × mass of one neutron

= 14×1 amu = 14 amu

Hence mass of the atom = 13 amu + 14 amu = 27 amu

ATOMIC MASS AS AN AVERAGE ATOMIC MASS

Atoms of the same element must have the same **number of protons** called **atomic number.**

- However, it is possible for two atoms of the same element to have different number of neutrons and hence difference in their masses. These *atoms of the same element with the same atomic number but differ in their masses* are known as **isotopes**.

Due to existence of isotopes, it is important to find the average mass in order to know mass of each atom which constitute to one particular element. For this reason, atomic mass is always given as an **average atomic mass**.

- As an example, consider an element whose atom, **X**, has two isotopes; one with mass of**35amu** and another with mass of **37amu**. If a sample of the element is taken such that number of atoms of **X – 35** isotope and **X – 37** isotopes in the sample are respectively **15** and **5**, then the average atomic mass may be found as follows:

From simple mathematics knowledge you studied even at primary level of education, it is clearly understood that:

Average atomic mass = $\dfrac{\textbf{total mass of all atomis in the sample}}{\textbf{total number of atoms in the sample}}$

But; total mass of all atoms in the sample

= total mass of atoms of **X-35**+ total mass of atoms of **X -37**

And total mass of **X-35** atoms

= number of **X-35** atoms× mass of one $\mathbf{X-35}$ atom

$= 15 \times 35\text{amu} = 525\ \text{amu};$

Total mass of **X-37** atoms

= number of **X-37** atoms ×mass of one $\mathbf{X-37}$ atom

$= 5 \times 37\text{amu} = 185\ \text{amu}$

It follows that;

Average atomic mass

$$= \frac{(\text{Number of X} - 35 \text{ atoms } \times \text{mass of one X} - 35) + (\text{number of X} - 37 \text{ atoms } \times \text{mass of one X} - 37 \text{ atom})}{\text{Total number of atoms in the sample}}$$

In a shorter way, the above expression may be rewritten as follows:

Average atomic mass

$$= \frac{\textbf{Sum of (number of atoms of particular isotope} \times \textbf{isotopic mass)}}{\textbf{Total number of atoms in the sample}}$$

$$= \frac{525\text{amu} + 185\text{amu}}{15 + 5} = \frac{710\text{amu}}{20} = 35.5 \text{ amu}$$

Have you noticed this ……………?

From the above example it is clearly understood that, the average atomic mass (35.5amu) is closer to the mass of **X-35** isotope which was more abundant than the mass of less abundant isotope, **X-37**.

- This is the general fact that, *average atomic mass is always closer to the atomic mass of an isotope with greater abundance in the isotopic mixture.*

Example 2

Boron exists in two isotopes: boron – 10 and boron -11. Atomic mass of boron is found to be 10.811amu. Based on this atomic mass, which isotope should be more abundant?

Answer

With given atomic mass of 10.811, boron -11 is more abundant because its mass number is closer to the atomic mass.

Understanding the formula for calculating average atomic mass

From our example in the previous discussion, we have seen that;

$$\textbf{Average atomic mass} = \frac{\textbf{total mass of all atoms}}{\textbf{total number of atoms}}$$

And we have also deduced that;

Total mass of all atoms = Sum of (number of atoms of particular isotope × isotopic mass)

And the final formula became;

$$\textbf{Average atomic mass } (\mathbf{A_r})$$
$$= \frac{\textbf{Sum of (number of atoms of particular isotope} \times \textbf{isotopic mass)}}{\textbf{Total number of atoms}}$$

Thus if the element has number of isotopes, say three isotopes with masses A_1, A_2 and A_3 and their respective number of atoms of each isotope n_1, n_2 and n_3, the above formula is then become:

$$\mathbf{A_r} = \frac{\mathbf{n_1A_1 + n_2A_2 + n_3A_3}}{\mathbf{n_1 + n_2 + n_3}}$$

If the abundances of isotopes are given as percentage which in actual sense they represent number of atoms of each isotope in 100 atoms of isotopic mixture, the above formula will be modified and become;

$$\mathbf{A_r} = \frac{\mathbf{P_1A_1 + P_2A_2 + P_3A_3}}{\mathbf{P_1 + P_2 + P_3}}$$

Where P_1, P_2 and P_3 are percentage abundances of each respective isotope.

But since summation of percentage is always 100, the above formula will become;

$\mathbf{A_r} = \frac{\mathbf{P_1A_1 + P_2A_2 + P_3A_3}}{\mathbf{100}}$ (For 100 atoms of the isotopic mixture)

In words, above formula may be rewritten as;

$$\mathbf{A_r} = \frac{\textbf{Sum of (percentage abundance} \times \textbf{isotopic mass)}}{\textbf{Total percentage}}$$

Similarly as the total percentage is always 100, the above formula become:

$$\mathbf{A_r} = \frac{\textbf{Sum of (percentage abundance} \times \textbf{isotopic mass)}}{\mathbf{100}}$$

Example 3

In a sample of 400 lithium atoms, it is found that 30 atoms are lithium−6 (6.015g/mol) and 370 atoms are lithium−7 (7.016g/mol).Calculate the average atomic mass of lithium.

Solution

Using$A_r = \frac{n_1A_1 + n_2A_2}{n_1 + n_2}$

Substituting $A_r = \left(\frac{370\times7.016+30\times6.015}{400}\right)$ g/mol = 6.941g/mol

Hence the average atomic mass of lithium is 6.941g/mol

Example 4

Naturally occurring europium (Eu) consists of two isotopes which have a mass of 151 and 153. Europium -151has an abundance of 48.03% and europium – 153 has an abundance of 51.97%. What is the atomic mass of europium?

Solution

Data given:

Europium-151 has percentage of 48.03%

Europium-153 has percentage of 51.97%

Required: atomic mass of europium, A_r

From the formula

$$A_r = \frac{\text{Sum of (isotopic mass}\times\text{percentage abundance)}}{\text{Total percentage}}$$

$$= \frac{(151 \times 48.03) + (153 \times 51.97)}{48.03+51.97}$$

$$= \frac{7252.53 + 7951.41}{100} = \frac{15{,}203.94}{100} = 152.04 \text{ amu}$$

Atomic mass of europium is 152.04 amu

Example 5

Titanium has five common isotopes:
$^{46}Ti(8.0\%)$, $^{47}Ti(7.8\%)$, $^{48}Ti(73.4\%)$, $^{49}Ti(5.5\%)$ and $^{50}Ti(5.3\%)$
What is the average atomic mass of titanium?

Solution

The superscript in the symbol of the atom element is the mass number which is approximately equal to the isotopic mass. Thus:

^{46}Ti: isotopic mass is 46; abundance 8.0%

^{47}Ti: isotopic mass is 47; abundance 7.8%

^{48}Ti: isotopic mass is 48; abundance 73.4%

^{49}Ti: isotopic mass is 49; abundance 5.5%

^{50}Ti: isotopic mass is 50; abundance 5.3%

Thus from the formula;

$$A_r = \frac{\text{Sum of (isotopic mass} \times \text{percentage abundance)}}{\text{Total percentage}}$$

$$= \frac{(46 \times 8) + (47 \times 7.8) + (48 \times 73.4) + (49 \times 5.5) + (50 \times 5.3)}{8 + 7.8 + 73.4 + 5.5 + 5.3}$$

$$= \frac{368 + 366.6 + 3523.2 + 269.5 + 265}{100} = 47.9\text{amu}$$

Average atomic mass of titanium is 47.9 amu

Example 6

The average atomic mass of copper is 63.55amu. If the only two isotopes of copper have masses of 62.94amu and 64.93amu. What are percentages of each?

Solution

Let P_1 be the percentage abundance of copper isotope, with mass of 64.93amu.

Then $P_2 = (100 - P_1)$ will be the percentage abundance of copper isotope with mass of 62.94amu. (Total percentage must be 100).

From the formula;

$$A_r = \frac{\text{Sum of (isotopic mass} \times \text{percentage abundance)}}{100}$$

$$= \frac{A_1P_1 + A_2P_2}{100}$$

$$A_r = \frac{62.94(100 - P_1) + (P_1 \times 64.93)}{100}$$

$$A_r = \frac{6294 - 62.94P_1 + 64.93P_1}{100}$$

But$A_r = 63.55$ (given):

$$63.55 = \frac{6294 + 1.99P_1}{100}$$

Thus $63.55 \times 100 = 6294 + 1.99P_1; 6355 = 1.99P_1 + 6294$

$1.99P_1 = 6355 - 6294; 1.99P_1 = 61; P_1 = \frac{61}{1.99} = 30.65$

$P_1 = 30.65\%$

But again; $P_2 = 100 - P_1$= 100 – 30.65 = 69.35%

Hence the percentages of the two isotopes are:

30.65% for copper atom with mass of 64.93 amu

69.35% for copper atom with mass of 62.94 amu.

Example 7

Naturally occurring silicon has three isotopes ^{28}Si , ^{29}Si and ^{30}Si whose atomic masses are 27.976amu, 28.9865amu and 29.9838amu respectively. The most abundant isotope is ^{28}Si which accounts for 92.23% of naturally occurring silicon. Given the observed atomic mass of silicon is 28.0855, calculate the percentage of ^{29}Si and ^{30}Si in nature.

Solution

Atomic mass of silicon, A_r = 28. 0855

Percentage of ^{28}Si $= 92.23\%$ with mass 27.976amu

Let the required percentage of ^{30}Si with mass 29.9838amu be P_1 Then the percentage of ^{29}Si with mass 28.9865amu will be

$P_2 = 100 - (P_1 + 92.23) = 7.77 - P_1$ (Summation of percentage must be 100)

Then from the formula;

$$A_r = \frac{\text{Sum of (isotopic mass} \times \text{percentage abundance)}}{100}$$

$$28.055 = \frac{(27.976 \times 92.23) + (P_1 \times 29.9838) + (7.77 - P_1)28.9865}{100}$$

$$2805.5 = 2580.23 + 225.23 - 28.9865P_1 + 29.9838P_1$$

$P_1 = \frac{2805.5-2805.46}{0.9965}$; $P_1 = 0.04\%$ and $P_2 = 7.77 - 0.04 = 7.73\%$

Hence the percentage of ^{30}Si and ^{29}Si are respectively 0.04% and 7.73%.

Example 8

Bromine has two naturally occurring isotopes; bromine – 79, has a mass of 78.91amu and 50.69% abundant. If the atomic mass of bromine is 79.904amu, determine the mass bromine – 81, the other isotope of bromine.

Solution

If the percentage abundant of bromine – 79 is 50.69%.

Then the percentage abundance of Br – 81 will be
$(100 - 50.69)\% = 49.31\%$.

Let the mass of bromine – 8 be M;

$$A_r = \frac{\text{Sum of (isotopic mass} \times \text{percentage abundance)}}{100}$$

$$A_r = \frac{(78.91 \times 50.69) + (M \times 49.31)}{100}$$

But $A_r = 79.904$; $79.904 = \frac{3999.95+49.31M}{100}$

$79.904 \times 100 = 3999.95 + 49.31M$

$7990.4 - 3999.95 = 49.31M$

$$3990.45 = 49.31M; M = \frac{3990.45}{49.31} = 80.926$$

Therefore, isotopic mass of bromine – 81 is 80.926amu.

Example 9

There are three isotopes of carbon. They have mass number of 12, 14, and 16. The average atomic mass of carbon is 12.0107.

(a) What does this say about the relative abundance of three isotopes?
(b) If the three isotopes listed in the question above all had the same relative abundance, what would the average atomic mass be?

Solution

(a) Carbon – 12 isotope is the most abundant because the average atomic mass is closer to 12 which is the mass number of carbon – 12.
(b) Given that: all isotopes have the same relative abundance. Total abundance = 100

Thus $100 = \%C - 12 + \%C - 14 + \%C - 16$

But $\%C - 12 = \%C - 14 = \%C - 16 = P$

Whence $100 = P + P + P;\ 100 = 3p;\ P = \frac{100}{3}$

$$A_r = \frac{\text{Sum of (isotopic mass} \times \text{percentage abundance)}}{\text{Total percentage}}$$

$$= \frac{\left(12 \times \frac{100}{3}\right) + \left(14 \times \frac{100}{3}\right) + \left(16 \times \frac{100}{3}\right)}{100} = 14\text{amu}$$

The average atomic mass would be 14amu

Alternative solution

If the isotopes have the same relative abundance, the average atomic mass becomes the arithmetic mean of given masses.

That is $A_r = \frac{A_1 + A_2 + A_3}{3} = \left(\frac{12+14+16}{3}\right)\text{amu} = \frac{42\text{amu}}{3} = 14\text{amu}$

Hence the average atomic mass is 14amu

PRACTICE EXERCISE 2

Question 1

Lithium – 6 is 4% abundant and lithium – 7 is 96% abundant. What is average mass of lithium?

Question 2

Iodine is 80% ^{127}I, 17% ^{126}I and 3% ^{128}I. Calculate the average atomic mass of iodine.

Question 3

What is the atomic mass of hafnium if, out of every 100 atoms, 5 have mass of 178, 14 have mass of 179 and 35 have a mass of 180?

Question 4

Calculate the percentage of each isotope present in the mixture of ^{113}In and ^{115}In which has an average mass of 14.8u.

Question 5

Calculate the atomic mass of silicon. The three silicon isotopes have atomic masses and relative abundances of 27.9769amu (92.2297%); 28.9765amu (4.6832%) and 29.9738amu (3.0872%).

Question 6

Antimony has two naturally occurring isotopes. The mass of antimony – 121 is 120.904amu and the mass of antimony 123 is 122.904amu. If the average atomic mass of antimony is 121.76amu, calculate the abundance of each isotope.

Question 7

Silicon has three stable isotopes: silicon – 28, atomic mass 27.98u and abundance 92.18%, silicon - 29 atomic mass 28.98u and abundance 4.71% and silicon – 30, atomic mass 29.97u. Before commencing the calculation and without a periodic table; what would you except approximately the average atomic mass of naturally occurring silicon be? Confirm your estimation by calculating the value.

Question 8

What is the atomic mass of osmium if out of every 100 atoms, 2 have a mass of 156amu, 12 have a mass of 157amu, 14 have mass of 158amu, 14 have a mass of 159amu and the remainder have a mass of 160amu?

Question 9

Calculate the percent abundance of the two isotopes of iridium, Ir – 191 and Ir – 193. The average atomic mass for iridium is 192.22amu.

Chapter three

MOLECULAR MASS AND PERCENTAGE COMPOSITION OF COMPOUNDS

INTRODUCTION

Molecular mass (or **molecular weight**) *is the mass of one molecule of an element or a compound.*

- It is the summation of mass of all atoms present in the molecule. Therefore, like atomic mass, molecular mass is given in atomic mass unit (amu).

However, the reader should understand that ***the term 'molecular mass' is applicable for covalent compounds only, why?***

- This is because, only covalent compounds exist as molecule. Since ionic compounds exist in network structure of oppositely charged ions and not as molecule, the term molecular mass is not applicable for ionic compounds.

Interesting ……….! If the term 'molecular mass' is not applicable for ionic compounds, what is the appropriate term for them?

- *Ionic compound exists as repeating unit of oppositely charged ions whereby each unit is known as* **formula unit**. For example; NaCl is the formula unit for sodium chloride and CaF_2 is the formula unit for calcium fluoride. Similarly for covalent compounds like phosphorous pentachloride, PCl_5 is its formula unit.
- *The mass of one formula unit is known as* **formula mass**. It is the summation of mass of all atoms present in one formula unit.

So formula mass is the appropriate term for ionic compounds although the term is applicable for covalent compounds too while the term molecular mass is applicable for covalent compounds only.

Example 1

Calculate molecular mass of the following:

(a) Cl_2
(b) CO_2

Solution

(a) Here there are two atoms of chlorine each with atomic mass of 35.5amu.

Then the total mass of atoms = $2 \times 35.5\text{amu} = 71\text{amu}$

Hence the molecular mass of $Cl_2 = 71\text{amu}$

(b) In the molecule of CO_2; there is one atom of C and two atoms of O.

But atomic mass of C = 12

And atomic mass of O = 16

Thus total mass of all atoms in the molecule

$$= \left(12 + (2 \times 16)\right)\text{amu} = 44\text{amu}$$

Hence molecular mass of CO_2 is 44amu.

Example 2

Calculate formula mass of the following:

(a) NaCl
(b) H_2SO_4
(c) $FeSO_4.7H_2O$

Solution

(a) Atomic mass of Na = 23amu

Atomic mass of Cl = 35.5amu

Total mass of all atoms in a unit of NaCl $= (23 + 35.5)\text{amu}$

$= 58.5\text{amu}$

Hence the formula mass of NaCl $= 58.5\text{amu}$

(b) Atomic mass of H = 1amu

Thus total mass of two H atoms in theH_2SO_4 unit

$= 2 \times 1\text{amu} = 2\text{amu}$

Atomic mass of S $= 32\text{amu}$

Atomic mass of O = 16amu

Thus total mass of four O atoms in H_2SO_4 unit

$= 4 \times 16\text{amu} = 64\text{amu}$

The total mass of all atoms in the formula unit

$= (2 + 32 + 64)\text{amu} = 98\text{amu}$

Hence formula mass of $H_2SO_4 = 98\text{amu}$

(c) Atomic mass of Fe $= 56$

Atomic of mass S = 32

Atomic mass of O =16

Atomic mass of H = 1

And in the one formula unit of $FeSO_4.7H_2O$ there are:

One Fe atom

One S atom
Eleven O atom and
Fourteen H atoms

Then the total mass of all atoms in the formula unit

$= \left(56 + 32 + (11 \times 16) + (14 \times 1)\right)$amu= 278amu

Hence the formula mass of$FeSO_4.7H_2O = 278\text{amu}$

Example 3

What is the molecular mass of a substance, each molecule of which contain 9 carbon atoms, 13 hydrogen atoms and 2.33×10^{-23}g of other component.

Solution

Mass of one carbon atom = 12 amu

Thus mass of carbon atoms = $12 \text{ amu} \times 9 = 108 \text{ amu}$

Mass hydrogen atom = 1 amu

Thus mass of 13 hydrogen atoms = $1 \text{ amu} \times 13 = 13 \text{ amu}$

Also $1.66 \times 10^{-24}\text{g} = 1 \text{ amu}$

Thus $2.33 \times 10^{-23}\text{g} = \frac{2.33 \times 10^{23}}{1.66 \times 10^{-24}} \text{ amu} = 14 \text{ amu}$

Then total mass of all atoms in one molecule

$$= (108 + 13 + 14) \text{ amu} = 135 \text{ amu}$$

Hence the molecular mass of the substance is 135 amu

PERCENTAGE COMPOSITION OF COMPOUNDS

Percentage composition *is the mass of each element in 100g of (or 100amu) of a compound.*

- It is the percentage by mass of each element in the compound.
- The summation of the percentages of all elements in the pure compound is always 100.

The percentage of particular element in a compound may be calculated by using the following formula;

$$\textbf{Mass percentage} = \frac{\textbf{mass of an element in a sample of a compound}}{\textbf{total mass of the compound in the sample}} \times \textbf{100}\%$$

If the sample is one molecule of the compound, the formula become:

$$\textbf{Mass percentage} = \frac{\textbf{mass of an element in 1 molecule of a compound}}{\textbf{molecular mass of the compound}} \times \mathbf{100\%}$$

If the sample is one mole of the compound, the formula become:

$$\textbf{Mass percentage} = \frac{\textbf{mass of an element in 1 mole of a compound}}{\textbf{molar mass of the compound}} \times \mathbf{100\%}$$

In calculating percentage composition:

- Firstly, calculate molecular mass (or formula mass) of the compound.
- Then divide the mass of each element in the one molecule of the compound by the obtained molecular mass (or formula mass).
- Finally, multiply each by 100%

Example 3

Calculate the percentage composition of CH_4

Solution

Molecular mass of $CH_4 = \left(12 + (4 \times 1)\right)\text{amu} = 16\text{amu}$

Mass of carbon in one molecule of $CH_4 = 12\text{amu}$

Thus % carbon $= \frac{\text{Mass of carbon in one molecule of } CH_4}{\text{Molecular mass of } CH_4} \times 100\%$

$$= \frac{12\text{amu}}{16\text{amu}} \times 100\% = 75\%$$

Mass of hydrogen in one molecular of $CH_4 = (4 \times 1)\text{amu} = 4\text{amu}$

Then% hydrogen $= \frac{\text{Mass of hydrogen in one molecule of } CH_4}{\text{Molecular mass of } CH_4} \times 100\%$

$$= \frac{4\text{amu}}{16\text{amu}} \times 100\% = 25\%$$

Hence the percentage composition of the compound is 75% carbon and 25% hydrogen

Example 4

What is the mass percentage of hydrogen and oxygen in water?

Solution

Molecular mass of $\mathrm{H_2O} = \big((2 \times 1) + 16\big)\mathrm{amu} = 18\mathrm{amu}$

$$\%\ \mathrm{Hydrogen} = \frac{\text{Mass of hydrogen in one molecule of } \mathrm{H_2O}}{\text{Molecular mass of } \mathrm{H_2O}} \times 100\%$$

$$= \frac{2 \times 1\mathrm{amu}}{18\mathrm{amu}} \times 100\% = \frac{2\mathrm{amu}}{18\mathrm{amu}} \times 100\% = 11.11\%$$

$$\%\ \mathrm{Oxygen} = \frac{\text{Mass of oxygen in one molecule of } \mathrm{H_2O}}{\text{Molecular mass of } \mathrm{H_2O}} \times 100\%$$

$$= \frac{16}{18} \times 100\% = 88.89\%$$

Hence:

Percentage composition of hydrogen = 11.11%

Percentage composition of oxygen = 88.89%

Example 5

Calculate the percentage composition of $\mathrm{C_6H_{12}O_6}$

Solution

Molecular mass of $\mathrm{C_6H_{12}O_6}$

$$= \big((6 \times 12) + (12 \times 1) + (6 \times 16)\big)\mathrm{amu} = 180\mathrm{amu}$$

$$\%\ \mathrm{Carbon} = \frac{\text{Mass of carbon}}{\text{Molecular mass}} \times 100\% = \frac{6 \times 12\mathrm{amu}}{180\mathrm{amu}} \times 100\% = 40\%$$

$$\%\ \mathrm{Hydrogen} = \frac{\text{Mass of hydrogen}}{\text{Molecular mass}} \times 100\%$$

$$= \frac{12 \times 1\mathrm{amu}}{180\mathrm{amu}} \times 100\% = 6.7\%$$

$$\%\ \mathrm{Oxygen} = \frac{\text{Mass of oxygen}}{\text{Molecular mass}} \times 100\% = \frac{6 \times 16\mathrm{amu}}{180\mathrm{amu}} \times 100\% = 53.3\%$$

Hence the percentage composition $C_6H_{12}O_6$is 40% carbon, 6.7% hydrogen and 53.3% oxygen.

Example 6

A 500mg tablet of aspirin, $C_9H_8O_4$, contains 300mg carbon and 8.08mg hydrogen, the remaining mass is oxygen. Determine percentage composition of aspirin.

Solution

Mass of aspirin = Mass of oxygen + Mass carbon + Mass of hydrogen

So500mg = Mass of oxygen + 300mg + 8.08mg

Thence mass of oxygen = 500mg – 308.08mg = 191.92mg

$$\text{Percentage composition} = \frac{\text{Mass element}}{\text{Total mass of compound}} \times 100\%$$

$$\text{Percentage of oxygen} = \frac{191.92\text{mg}}{500\text{mg}} \times 100\% = 38.4\%$$

Percentage of carbon = $\frac{300\text{mg}}{500\text{mg}} \times 100\%$ = 60%

$$\text{Percentge of hydrogen} = \frac{8.08\text{mg}}{500\text{mg}} \times 100\% = 1.6\%$$

Percentage composition is 38.4% oxygen, 60% of carbon and 1.6% of hydrogen.

Example 7

Calculate mass of carbon in grams, contained in 1kg of glucose,$C_6H_{12}O_6$.

Solution

Given: 1kg= 1000gof glucose

Molecular mass of glucose, $C_6H_{12}O_6$;

$= (12 \times 6) + (1 \times 12) + (16 \times 6)$amu= 180amu

Mass of carbon in $C_6H_{12}O_6 = (12 \times 6)$amu= 72amu

Percentage of carbon = $\frac{72\text{amu}}{180\text{amu}} \times 100 = 40\%$

Total mass of glucose = 1kg = 1000g

So mass of carbon in 1000g

Mass of carbon = % of carbon by mass × Total mass of compound

$= \frac{40}{100} \times 1000\text{g} = 400\text{g}$

Mass of carbon is 400g

Example 8

Calculate the percentage of nitrogen in ammonium sulphate.

Solution

Molecular mass of compound,$(NH_4)_2SO_4$;

$$= \left(14 + (1 \times 4)\right) \times 2 + 32 + (16 \times 4) = 132\text{amu}$$

Mass of nitrogen in $(NH_4)_2SO_4$= $(14 \times 2)\text{amu} = 28\text{amu}$

Percentage of nitrogen = $\frac{28\text{amu}}{132\text{amu}} \times 100\% = 21.21\%$

Percentage of nitrogen is 21.21%

Example 9

Find the percentage composition of a pure substance that contains 7.22g nickel, 2.53g phosphorous and 5.25g oxygen only.

Solution

Total mass $= 7.22\text{g} + 2.53\text{g} + 5.25\text{g} = 15\text{g}$

Percentage of Nickel = $\frac{\text{Mass of Nickel}}{\text{Total mass of the compound}} \times 100\%$

$= \frac{7.22g}{15g} \times 100\% = 48.1\%$

Percentage of phosphorous = $\frac{\text{Mass phosphorous}}{\text{Total mass of the compound}} \times 100\%$

$$= \frac{2.535}{15g} \times 100\% = 16.9\%$$

Percentage of oxygen = $\frac{\text{Mass of oxygen}}{\text{Total mass of the compound}} \times 100\%$

$$= \frac{5.25g}{15g} \times 100\% = 35\%$$

Percentage composition is 48.1% Nickel, 16.9% phosphorous and 35% oxygen.

Example 10

A sample Ag_2S has a mass of 62.4g. What mass of each element could be obtained by decomposing this sample?

Solution

Finding composition of each element in the sample:

$$\text{Composition of Ag} = \frac{\text{Mass of Ag in } Ag_2S}{\text{Molecular mass of } Ag_2S} \times 100\%$$

$$= \frac{108 \times 2\text{amu}}{((108 \times 2) + 32)\text{amu}} \times 100\% = 87.1\%$$

$$\text{Composition of S} = \frac{\text{Mass of S in } Ag_2S}{\text{Molecular mass of } Ag_2S} \times 100\%$$

$$= \frac{32\text{amu}}{((108 \times 2) + 32)\text{amu}} \times 100\% = 12.9\%$$

So, mass of Ag in Ag_2S = Composition Ag × Total mass of Ag_2S

$$= \frac{87.1}{100} \times 62.4g = 54.4g$$

Mass of S in Ag_2S = Composition of S × Mass of Ag_2S

$$= \frac{12.9}{100} \times 62.4\text{g} = 8\text{g}$$

Hence mass of Ag will be 54.4g and sulphur will be 8g.

Example 11

Calculate the mass of zinc in a 30g sample of zinc nitrate.

Solution

Firstly, finding percentage composition of zinc in zinc nitrate:

Molecular mass zinc nitrate, $Zn(NO_3)_2$,

$$= (65 + (14 + (16 \times 3)) \times 2)\text{amu} = (65 + 124)\text{amu} = 189\text{amu}$$

$$\text{Then percentage Composition of Zn} = \frac{\text{Mass Zn in } Zn(NO_3)_2}{\text{Molecular mass}} \times 100\%$$

$$= \frac{65\text{amu}}{189\text{amu}} \times 100\% = 34.4\%$$

$$\text{Mass of zinc} = \text{Percentage compoition of Zn} \times \text{Mass of zinc nitrate}$$

$$= \frac{34.4}{100} \times 30\text{g} = 10.32\text{g}$$

Hence the mass of zinc is 10.32g

Example 12

A quantity of Epsom salts, magnesium sulphate heptahydrate, $MgSO_4.7H_2O$, is heated until all the water is driven off. The sample loses 11.8g in the process. What was the mass of original sample?

Solution

Given: mass loss in the heating process = mass of $H_2O = 11.8\text{g}$

Firstly finding the percentage composition of water in $MgSO_4.7H_2O$

Molar mass of $MgSO_4.7H_2O$

$$= 24 + 32 + (16 \times 4) + 7((1 \times 2) + 16) = 246\text{g/mol}$$

Mass of H_2O in $MgSO_4.7H_2O$

$= 7 \times ((1 \times 2) + 16) = 7 \times 18 = 126g/mol$

Percentage of $H_2O = \frac{126/mol}{246/mol} \times 100\% = 51.2\%$

Mass of H_2O = Percentage of water $\times$ Total mass

Hence total mass = $\frac{\text{Mass of } H_2O}{\text{Percentage of } H_2O} = \frac{11.82}{51.2/_{100}}$

$$= \left(\frac{11.8 \times 100}{51.2}\right) g \quad = 23g$$

Mass of the original sample is 23g

Example 13

Magnesium is the metallic element present in chlorophyll.

Analysis of chlorophyll revealed that it contains 0.04% of metal. Determine the minimum possible molar mass of chlorophyll.

Solution

Minimum molar mass is obtained if there is only one Mg atom per one molecule of the chlorophyll.

That is %Mg = $\frac{\text{Molar mass of Mg atoms}}{\text{Molar mass of chlorophyl}} \times 100\%$

Substituting 0.04 = $\frac{24}{\text{Molar mass of chlorophyll}} \times 100$

From which molar mass of chlorophyll = $\frac{24 \times 100}{0.04} = 60000g/mol$

Hence the minimum possible molar mass of chlorophyll is 60000g/mol

Example 14

If a pure compound is composed of X_2Y_3 molecules and consists of 60% X by weight. What is the atomic weight of Y in terms of atomic weight of X?

Solution

Let atomic weight of X and Y be $M_x$and M_y respectively.

Then %X = $\frac{\text{Total mass of X in one molecule of } X_2Y_3}{\text{Molecular mass of } X_2Y_3} \times 100\%$

Where molecular mass of $X_2Y_3 = 2M_x + 3M_y$

If follows that; $60 = \frac{2M_x}{2M_x+3M_y} \times 100$

$0.6(2M_x + 3M_y) = 2M_x$; $1.2M_x + 1.8M_y = 2M_x$

$1.8\ M_y = 0.8M_x$; $M_y = \frac{0.8}{1.8}M_x = \frac{4}{9}M_x$

Hence $M_y = \frac{4}{9}M_x$

APPLICATION OF PERCENTAGE COMPOSITION IN CALCULATIONS INVOLVING SOIL CHEMISTRY

Example 15

A certain soil requires 80kg of nitrogen (N) per hectare so as to fulfill plant requirements of N. Calculate (in kg) the quantity of ammonia sulphate ($(NH_4)_2SO_4$) fertilizer required to meet this demand.

Solution

Molar mass of $(NH_4)_2SO_4$

$$= \left((2 \times 18) + 32 + (4 \times 16)\right)\text{g/mol} = 132\text{g/mol}$$

Mass of N atoms in one mole of $(NH_4)_2SO_4 = 2 \times 14 = 28\text{g}$

$$\%N = \frac{\text{Mass of N in one mole of}(NH_4)_2SO_4}{\text{Molar mass of }(NH_4)_2SO_4} \times 100\%$$

$$= \frac{28}{132} \times 100\% = 21.21\%$$

Then using; mass of nitrogen = % of N × total mass of $(NH_4)_2SO_4$;

Substituting $80\text{kg} = \frac{21.21}{100} \times \text{Total mass of } (NH_4)_2SO_4$

From which; total mass of $(NH_4)_2SO_4 = \frac{100 \times 80\text{kg}}{21.21} = 377\text{kg}$

Hence 377kg of ammonium sulphate fertilizer will be required.

Example 16

A farmer was advised to supply 160kg of N to his maize farm. Calculate the mass of fertilizer which has to buy to meet the requirements.

Given that: the fertilizer is 80% by mass$Ca(NO_3)_2$.

Solution

Molar mass of $Ca(NO_3)_2$

$$= \big(40 + (2 \times 14) + (6 \times 16)\big)\text{g/mol} = 164\text{g/mol}$$

Mass of N atoms in one mole of $Ca(NO_3)_2 = 2 \times 14 = 28\text{g}$

$$\%N = \frac{\text{Mass of N in one mole of } Ca(NO_3)_2}{\text{Molar mass of } Ca(NO_3)_2} \times 100\%$$

$$= \frac{28}{164} \times 100\% = 17.07\%$$

Then using; mass of nitrogen = % of N × Total mass of $Ca(NO_3)_2$;

Substituting $160\text{kg} = \frac{17.07}{100} \times \text{Total mass of } Ca(NO_3)_2$

From which; total mass of $Ca(NO_3)_2 = \frac{100 \times 160\text{kg}}{17.07} = 937\text{kg}$

Thus 937kg of $Ca(NO_3)_2$will be required.

But the fertilizer is 80%$Ca(NO_3)_2$

Thus using $\%Ca(NO_3)_2 = \frac{\text{mass of } Ca(NO_3)_2}{\text{total mass of fertilizer}} \times 100$

$80 = \frac{937}{\text{total mass of fertilizer}} \times 100;$

From which total mass of fertilizer $= \frac{937 \times 100}{80}\text{kg} = 1171\text{kg}$

Hence 1171kg of the fertilizer is required to meet the requirement.

PRACTICE EXERCISE 3

Question 1

Calculate the following:

(a) Percentage of potassium in potassium phosphate.
(b) Percentage of nitrogen in ammonium nitrate.
(c) Mass of carbon in 16g of methane.
(d) Percentage composition of phosphorous in calcium phosphate.

Question 2

A sample of compound is analyzed and found to contain 0.9g of calcium ad 1.6g of chlorine. The sample has a mass of 2.5g. Find the percentage composition of the compound.

Question 3

A sample of a compound is analyzed and found to contain carbon, hydrogen and oxygen. The mass of sample is 650mg and sample contains 257mg: of carbon and 50.4mg of hydrogen, what is the percentage composition of the compound.

Question 4

Sulphuric acid, H_2SO_4 is an important acid in laboratories and industries. Determine the percentage composition of sulphuric acid.

Question 5

Determine the percentage composition of sodium carbonate.

Question 6

Determine the percentage composition of each of the following compound:

(a) Sodium oxalate ($Na_2C_2O_4$)
(b) Ethanol (C_2H_5OH)

(c) Aluminium oxide
(d) Potassium sulphate

Question 7

Calculate the mass of given element in each of the following compound:

(a) Bromine in 50g of potassium bromide.
(b) Chromium in 1kg sodium dichromate,$Na_2Cr_2O_7$.
(c) Nitrogen in 85mg of amino acid lysine,$C_6H_{14}N_2O_2$.
(d) Cobalt in 2.84g of cobalt (II) acetate $Co\ (C_2H_8O_2)$

Question 8

The process of manufacturing sulphuric acid begin with the burning sulphur. What mass of sulphur would have to be burned in order to produce 1kg of H_2SO_4? Assume that all the suphuric end up in sulphuric acid.

Question 9

A sample of protein was analyzed for metal content analysis revealed that it contained magnesium and titanium in equal amount (by weight). If these are the only metallic species present in the protein and it contains 0.015% metals by weight. Determine minimum possible molar mass of this protein.

(Atomic masses: $Mg = 24, Ti = 48$)

Chapter four

THE MOLE

INTRODUCTION

A **mole** is a unit which is used to express the amount of substances.

- It can simply be defined as *the amount of substance of which contains Avogadro's number of particles.*
  Whereby: Avogadro's number, $N_A = 6.02 \times 10^{23}$

Thus one mole of any substance contains 6.02×10^{23} particles.

- The particles may be either atoms, ions, molecules, electrons, protons or neutrons.

Number of moles of substances can be calculated from one of the following:

(i) Counted number of particles in the substance.
(ii) Measured mass of the substance
(iii) Measured volume of the gaseous substance
(iv) Determined molarity of the solution(We will have detailed discussion on this in the **chapter five**).

NUMBER OF MOLES FROM NUMBER OF PARTICLES

If:

N represents number of particles contained in the substance

$\mathbf{N_A}$ takes its usual meaning of Avogadro's number (constant) and

n also takes its usual meaning of number of moles.

Then from the definition of mole: **1**mole $\equiv$ $\mathbf{N_A}$ particles

And for **N** particles corresponding to **n** moles: **n** mole $\equiv$ **N** paticles

Then by cross multiplication;

$$\begin{matrix} 1 \equiv N_A \\ n \equiv N \end{matrix}$$

$\mathrm{n} \times \mathrm{N_A} = 1 \times \mathrm{N}$

From which; $\mathrm{n} = \frac{\mathrm{N}}{\mathrm{N_A}}$ (Dividing by $\mathrm{N_A}$ both sides)

Hence number of moles of a substance from the given number of particles in the substance may be calculated by the following formula:

$$\mathbf{n} = \frac{\mathbf{N}}{\mathbf{N_A}}$$

Where:

n is the number of moles

N is the number of particles (atoms, ions, molecules, electrons, protons or neutrons)

$\mathrm{N_A}$ is the Avogadro's number = 6.02×10^{23}(particles/mol)

NUMBER OF MOLES FROM MEASURED MASS

Mass of one mole of a substance is known as **molar mass**, $\mathbf{M_r}$**.**Its common unit is g/mol ($\mathrm{gmol^{-1}}$)

- So in other words, **molar mass ($\mathbf{M_r}$)** *is the total mass of Avogadro's number* (6.02×10^{23}) *of particles contained in the substance.*

That is;

$\mathbf{M_r}$= Mass of one mole mass = Total mass of $\mathbf{6.02 \times 10^{23}}$ particles

Now, if we have measured mass of any substance, say **m**, number of moles of the substance can be found as follows:

From the definition of molar mass: $\mathbf{1}$ mole $\equiv \mathbf{M_r}$

And for measured mass, **m**, corresponding to **n** moles; $\mathbf{n}$ mole $\equiv \mathbf{m}$

Then by cross multiplication;

$$\begin{array}{ccc} 1 & \equiv & M_r \\ n & \equiv & m \end{array}$$

$n \times M_r = 1 \times m$

From which; $n = \frac{m}{M_r}$ (Dividing by M_r both sides)

Hence number of moles of the substance from the given mass of the substance may be calculated by the following formula:

$$\mathbf{n = \frac{m}{M_r}}$$

Where:

n is the number of moles as usual

m is the mass of the substance (commonly in grams, g)

$\mathbf{M_r}$ is the molar mass of the substance in the same unit (of mass) as of m (commonly g/mol)

Molar mass from number of particles

If **N** particles have total mass of **m**, then:

$n = \frac{N}{N_A}$ and $n = \frac{m}{M_r}$;

It follows that (by equating);

$$\frac{m}{M_r} = \frac{N}{N_A}$$

From which (by making M_r the subject);

$$\mathbf{M_r = \frac{mN_A}{N}}$$

But $\frac{m}{N}$ = Mass of one particle;

Hence $\mathbf{M_r} = \mathbf{Mass\ of\ one\ particle} \times \mathbf{N_A}$

Interesting relationship between amu and g/mol

You should be able to remember that in the **chapter two** we noted that $\mathbf{1amu = 1.66 \times 10^{-24}g}$.

- Now, assume we have one particle of a substance with a certain mass in amu say **x**.

That is mass of one particle of the substance = **x** amu

But because 1 amu $= 1.66 \times 10^{-24}$g

Then **x** amu $= 1.66 \times 10^{-24}\mathbf{x}$ g

- Thus mass of one particle will be $1.66 \times 10^{-24}\mathbf{x}$g
- So if we have Avogadro's number,6.02×10^{23} (**one mole**) of particle, the total mass will then become:

$$1.66 \times 10^{-24}\mathbf{x}\ ^{g}/_{particle} \times 6.02 \times 10^{23}\ ^{particle}/_{mol} = \mathbf{x}\ g/mol$$

- That is if **x** amu is the mass of **one particle**, then the mass of **one mole** of particles will be **x** g.

Thus if the mass of the **one particle** is 1 amu, mass of **one mole of particles** will be 1g/mol. This leads us to very important, and interesting conclusion on the relationship between amu and g/mol which is:

Numerical value of mass of one particle of a substance in amu is always equal to the numerical value of mass of one mole of particles of the same substance in in grams (g).

The above conclusion in the bolded italic, also leads us to the following useful facts:

- Numerical value of atomic mass (in amu) is equal to the numerical value of molar mass of the atom (in g/mol). For example, atomic mass of sodium is **23**amu and molar mass of the sodium is also **23**g/mol.

- Numerical value of molecular mass (or formula mass) of the compound (in amu) is equal to the numerical value of molar mass of the compound (in g/mol). For example, molecular mass of $\mathrm{CH_4}$ is **16**amu while molar mass of $\mathrm{CH_4}$ is also **16** g/mol.

The above two 'dots' explain why **amu and g/mol are commonly used interchangeably for atomic mass and molecular mass.**

NUMBER OF MOLES FROM MEASURED VOLUME

For gaseous (not liquid or solid), volume is directly proportional to the number of moles of the substance (provided that temperature and pressure is kept constant).

- Volume of one mole of the gaseous substance at given temperature and pressure is known as m**olar volume of a gas or gram molecular volume (GMV).** Its value is the same for all gases provided that temperature and pressure remain constant.

So in other words, **GMV** *is the total volume of Avogadro's number* (6.02×10^{23}) *of molecules (particles) of gaseous substance (at given temperature and pressure).*

That is:

$\mathrm{GMV} = \text{Volume of one mole} = \text{Total volume of } 6.02 \times 10^{23} \text{ gas molecules}$Now; if we have measured volume of any gaseous substance, say V, number of moles of the substance can be found as follows:

From the definition of Gas Molar Volume (GMV): **1** mole $\equiv$ **GMV**

And for measured volume, **V**, corresponding to **n**moles: **n** mole $\equiv$ **V**

Then by cross – multiplication:

$$\begin{matrix} 1 \equiv \mathrm{GMV} \\ n \equiv V \end{matrix}$$

$n \times GMV = V$

From which; $n = \frac{V}{GMV}$(Dividing by GMV both sides)

Hence number of moles of the substance from the given volume of the substance may be calculated by the following formula:

$$\mathbf{n = \frac{V}{GMV}}$$

Where:

V is the volume of the gaseous substance at given temperature and pressure.

GMV is the gas molar volume of the substance at the same temperature and pressure as that involved in taking measurement of V.

Be careful:

- The above formula is applicable if and only if the substance is gas. It is clear mistake to apply the formula in liquid and solid substance.
- When applying the formula, make sure that temperature and pressure for measured GMV is the same as that for measured V.

Also the reader should understand that, at standard temperature and pressure (STP) which is the temperature of 0°C (273K) and pressure of 1atm (760mmHg or 101325Nm^{-2}), GMV for all gases is 22.4dm^3 (22.4 L).

- So if measured volume is taken at STP, the above formula can be modified to:

$$\mathbf{n = \frac{V}{22.4\ dm^3}}$$

Where V must be in dm^3 (the same unit as GMV).

Molar mass from measured volume of gas

If **V** is the volume of a gas with mass of **m**, then:

$n = \frac{V}{GMV}$ and $n = \frac{m}{M_r}$;

It follows that (by equating);

$$\frac{m}{M_r} = \frac{V}{GMV}$$

From which (by making M_r the subject);

$$\mathbf{M_r = \frac{m \times GMV}{V}}$$

But $\frac{m}{V}$ = Density of a gas in gdm^{-3};

Hence $\mathbf{M_r = Density\ of\ a\ gas \times GMV}$

If M_r is in $gmol^{-1}$ and

GMV is in dm^3mol^{-1}

Density of a gas must be in gdm^{-3}

Example 1

Calculate the number of moles in:

(i) 60g of calcium
(ii) An iron sample containing 10^{22} atoms of iron

Solution

(i) Mass of calcium = 60 g

Molar mass of calcium = 40 g/mol

Number of moles;$n = \frac{m}{M_r} = \frac{60g}{40g/mol} = 1.5mol$

(ii) Number of atoms = 10^{22} atoms

Thenn $= \frac{N}{N_A} = \frac{10^{22}}{6.02 \times 10^{23}} = 0.0166$ mol

Example 2

Calculate the molar mass of each of the substances mentioned in the following:

(a) A 0.00496 mol sample of cholesterol has a mass of 1.894g
(b) The mass of a 3.44×10^{-5}mol sample of a particular protein has a mass of 74.8 g

Solution

(a) From $n = \frac{m}{M_r}$; where, n = 0.00496 mol, m = 1.894 g

Substituting $0.00496\text{mol} = \frac{1.894\ g}{M_r}$

$M_r = \frac{1.894\ g}{0.00496\ mol} = 381.85$ g / mol

(b) Number of mole, $n = 3.44 \times 10^{-5}$ mol ; m = 74.8 g

From $n = \frac{m}{M_r}$; $3.44 \times 10^{-5}\ mol = \frac{74.8\ g}{M_r}$

From which $M_r = \frac{74.8g}{3.44 \times 10^{-5}\ mol} = 2.17 \times 10^6$ g/mol

Example 3

Calculate volume of STP occupied by the following:

(a) 24.8mol of NH_3
(b) 3.66g of HCl

Solution

(a) Number of mole, n = 24.8 mol:

Using, $n = \frac{V}{GMV}$; where GMV at STP = 22.4 $^{dm^3}/_{mol}$

From which $V = n \times GMV$

$$= 24.8 \text{ mol } \times 22.4 \ ^{\text{dm}^3}/_{\text{mol}} = 555.52 \text{ dm}^3$$

(b) Given mass of HCl $= 3.65$ g

From $n = \frac{m}{M_r}$

Also $n = \frac{V}{GMV}$

Then by equating $n = \frac{m}{M_r} = \frac{V}{GMV}$

From which (by making V the subjects) $V = \frac{G.M.V \times m}{M_r}$

Where M_r of HCl $= 35.5+1 = 36.5$; GMV $= 22.4 \text{ dm}^3$

Thus $V = \dfrac{22.4 \times 3.65}{36.5} = 2.24 \text{ dm}^3$

Hence the volume of HCl is 2.24 dm^3

Example 4

What volume at STP occupied by each of the following:

(a) 8.27×10^{20} molecules of O_2
(b) 0.725mol of carbon dioxide gas

Solution

(a) Given 8.27×10^{20} molecules of O_2

Using $n = \frac{N}{N_A} = \frac{V}{GMV}$

Substituting $\frac{8.27 \times 10^{20}}{6.02 \times 10^{23}} = \frac{V}{22.4}$

From which $V = \frac{8.27 \times 10^{20}}{6.02 \times 10^{23}} \times 22.4 \text{ dm}^3 = 0.03077 \text{ dm}^3$

Hence the volume is 0.03078 dm^3 or 30.77cm^3

(b) Using $n = \frac{V}{GMV}$;

From which $V = n \times GMV$

$$= 0.725\text{mol} \times 22.4\text{dm}^3/\text{mol} = 16.24\text{dm}^3$$

Hence the volume is 16.24dm^3

Example 5

Calculate:

(a) Number of moles of atoms present in 88g of CO_2
(b) Mass of 2mol of H_3PO_4

Solution

(a) Given mass of CO_2, $m = 88\text{g}$

Using $n = \frac{m}{M_r}$; Where M_r of $CO_2 = 12 + (2 \times 6) = 44\ \text{g/mol}$

$$n = \frac{88\ \text{g}}{44\ \text{g / mol}} = 2\ \text{mol}$$

But one mole of CO_2 contains one mole of carbon atom and two moles of oxygen atoms making a total of three moles of atoms per one mole of CO_2.

Thus total number of moles atoms in 2moles of CO_2 is;

$$3 \times 2\text{moles} = 6\text{moles of atoms}$$

(b) Given number of moles,$n = 2\text{mol}$

$n = \frac{m}{M_r}$; $m = n \times M_r$

WhereM_r of $H_3PO_4 = (1 \times 3) + 31 + (16 \times 4) = 98\ \text{g / mol}$

Thus $m = 2\ \text{mol} \times 98\ \text{g} = 196\ \text{g}$

Hence mass of 2mol of H_3PO_4 is $196\ \text{g}$

Example 6

Calculate the number of:

(a) Molecules in 0.325mol of F_2
(b) Formula units in 0.75mol of CaF_2

Solution

(a) Given number of mole of $F_2, n = 0.325$ mol

So $N = nN_A$

$= 0.325 \times 6.02 \times 10^{23} = 1.9565 \times 10^{23}$ molecules

(b) Given 0.75mol of CaF_2

Number of formula units, $N = n\, N_A$

$= 0.75 \times 6.02 \times 10^{23} = 4.515 \times 10^{23}$ formula units

Therefore there are 4.52×10^{23} formula units

Example 7

How many atoms of sodium are in 13g of sodium metal?

Solution

Molar mass of sodium = 23 g mol $^{-1}$

From $n = \frac{N}{N_A}$; $N = nN_A$

But $n = \frac{m}{M_r}$; where mass of sodium = 13g

$$\text{Thus } N = \frac{m}{M_r}N_A = \frac{13 \times 6.02 \times 10^{23}}{23} = 3.403 \times 10^{23}$$

Hence the total number of atoms = 3.403×10^{23}atoms

Example 8

A sample of nitrogen contains 5.6×10^{19} atoms of nitrogen. Find the mass of atoms.

Solution

Total number of atoms = 5.6×10^{19} atoms

Usingn $= \frac{N}{N_A} = \frac{m}{M_r}$; where M_r of nitrogen atoms $= 14$ g/mol

Substituting $\frac{5.6 \times 10^{19}}{6.02 \times 10^{23}} = \frac{m}{14}$

From which m $= \frac{5.6 \times 10^{19}}{6.02 \times 10^{23}} \times 14\text{g} = 1.3 \times 10^{-3}$ g

Hence the mass of oxygen atoms is 1.3×10^{-3} g

Example 9

Determine the number of sulphur atoms in 1.59mmol of carbon disulphide.

Solution

Chemical formula of carbon disulphide is C S_2

Number of mole of CS_2=1.59mmol(millimoles)

=1.59×10^{-3} mol

From the formula: CS_2: 1 mole of $CS_2 \equiv 2$ mol of S

Thus 1.59×10^{-3} mol of $CS_2 \equiv 1.59 \times 10^{-3} \times 2$mol

$= 3.18 \times 10^{-3}$ mol

Then using $N = nN_A$;

Number of sulphur atoms, N

$= 3.18 \times 10^{-3} \times 6.02 \times 10^{23}\ {}^{\text{atom}}/_{\text{mol}} = 1.914 \times 10^{21}$ atom

Hence there are1.914×10^{21} sulphur atoms

Example 10

How many atoms of oxygen are in 1.5 mole of aluminium sulphate?

Solution

From $n = \frac{N}{N_A}$; $N = nN_A$

Number of moles of aluminium sulphate = 1.5mol

$$N_A = 6.02 \times 10^{23}\ \text{molecules}/_{\text{mol}}$$

$$N = 1.5\text{mol} \times 6.02 \times 10^{23}\ \text{molecules}/_{\text{mol}}$$

$$= 9.03 \times 10^{23}\ \text{molecules of } Al_2(SO_4)_3$$

Also chemical formula of aluminium sulphate is$Al_2(SO_4)_3$

From which number of O atoms in one molecule of $Al_2(SO_4)_3$ is 12.

Thence number of O atoms in 9.03×10^{23} molecules of $Al_2(SO_4)_3$ is$12 \times 9.03 \times 10^{23}$atoms $= 1.0836 \times 10^{25}$atoms.

Hence there are 1.0836×10^{24}atoms of oxygen

Example 11

In 3 moles of ethane, calculate the:

(i) Number of moles of carbon atoms.
(ii) Number of moles of hydrogen atoms.
(iii) Number of molecules of ethane.

Solution

(i) Chemical formula of ethane isCH_3CH_3
1 mole of ethane ≡ 2 mol of carbon
3 mole of ethane≡ $\frac{3 \times 2}{1}$= 6 mol of carbon
Hence there are 6 moles of carbon

(ii) 1 mol of ethane = 6 mol of hydrogen atoms

3 mol of ethane = $\frac{3 \times 6}{1}$ = 18 mol of hydrogen atoms

Hence there are 18 mol of hydrogen atoms

(iii) From $n = \frac{N}{N_A}$; $N = n\,N_A$

Where n = 3mol (given), $N_A = 6.02 \times 10^{23}$ (constant)

Substituting $N = 3 \times 6.02 \times 10^{23}$

$= 1.806 \times 10^{24}$ molecules

Hence there are 1.806×10^{24} molecules

Example 12

How many atoms are contained in the following?

(a) 2240mL of Cl_2
(b) 15 molecules of NH_4Cl
(c) 2.56 moles of $(NH_4)_3PO_4$
(d) 25g of $CuSO_4.5H_2O$

Solution

(a) Given 2240mL of Cl_2 = 2240cm^3 of Cl_2 = 2.24dm^3

From N = n NA; where $n = \frac{V}{GMV} = \frac{2.24\ dm^3}{22.4\ dm^3 mol^{-1}} = 0.1$ mol

Then $N = 0.1 \times 6.02 \times 10^{23}$

$= 6.02 \times 10^{22}$ molecules of Cl_2

But each molecule of Cl_2 contain two atoms of Cl

Thus total number of atoms = $2 \times 6.02 \times 10^{22}$ atoms

$= 1.204 \times 10^{23}$ atoms

Hence there are 1.204×10^{23} atoms

(b) 1 molecule of NH_4Cl

= 1 atom of N + 4 atom of H + 1 atom of Cl = 6 atoms

Thus1 molecule = 6 atoms

Then 15 molecule = $\frac{15 \times 6}{1}$ atoms= 90atoms

Hence there are 90 atoms

(c) From $\mathrm{n} = \frac{\mathrm{N}}{\mathrm{N_A}}; \mathrm{N} = \mathrm{n\ N_A}$

$= 2.56 \times 6.02 \times 10^{23} = 1.5411 \times 10^{24}$ molecules

But each molecule of $(NH_4)_3PO_4$ contains three N atoms, twelve H atoms, one P atom and four O atoms making a total of 20 atoms.

Thus the total number of atoms

$= 20 \times 1.5411 \times 10^{24}$ atoms $= 3.0822 \times 10^{25}$ atoms.

(d) Molar mass of $CuSO_4.5H_2O$

$= \left(64 \times 32 + (16 \times 4)\right) + (5 \times [(1 \times 2) + 16])$

$= 250\mathrm{g\ /mol}$

Then substituting $\mathrm{n} = \frac{\mathrm{m}}{\mathrm{M_r}} = \frac{25\mathrm{g}}{250\ \mathrm{g\ mol^{-1}}} = 0.1\mathrm{mol}$

But in one mole of $CuSO_4 . 5H_2O$, there is 1 mole of Cu atoms, 1 mol of S atoms, 9mol of O atoms and 10mol of H atoms making a total of 21mol of atoms.

Thus total number of moles of atoms in 1mol of $CuSO_4 . 5H_2O$, is 21mol.

And therefore in 0.1 mol of $CuSO_4 . 5H_2O$, number of moles will be $0.1 \times 21\mathrm{mol} = 2.1\mathrm{mol}$

Hence number of atoms, $\mathrm{N} = \mathrm{n\ N_A}$

$= 2.1 \times 6.02 \times 10^{23}$ atoms $= 1.2642 \times 10^{24}$ atoms

Example 13

A piece of copper contains 6.02×10^{24} atoms. How many mole of copper does it contain?

Solution

Given number of copper atoms, $N = 6.02 \times 10^{24}$ atom

Using $n = \frac{N}{N_A}$ where $N_A = 6.02 \times 10^{23}\ \text{atom}/_{\text{mol}}$

$n = \frac{6.02 \times 10^{24}}{6.02 \times 10^{23}\text{atom}/_{\text{mol}}}$=10mol

Hence the number of moles of copper $= 10$ mol

Example 14

Calculate:

(a) Number of moles of 2cm^3 of liquid bromine
(b) Number of Br atoms in$2\text{cm}^3$of liquid bromine

Given that density of liquid bromine is 3.1gcm^{-3}

Solution

(a) Given : Volume of liquid bromine = 2 cm^3
Density of liquid bromine = 3.1gcm^{-3}

Using; $\text{mass, m} = \text{Density, (p)} \times \text{Volume (v)}$

$\text{Mass of bromine} = 3.1\text{ g cm}^3 \times 2\text{ cm}^3 = 3.1\text{ g} \times 2 = 6.2\text{ g}$

Thus mass of bromine is 6.2 g

$\text{Thenn} = \frac{m}{M_r}$; where $M_r(Br_2) = 2 \times 80\text{ g / mol} = 160\text{g / mol}$

$$n = \frac{6.2}{160\text{ g / mol}} = 0.03875\text{ mol}$$

Hence number of moles of 2cm^3 of liquid bromine is 0.03875mol.

(b) In one mole of Br_2 there are 2moles of Br atoms.

Thus number of moles of Br atoms in 0.03875 mol of Br_2

$$= 2 \times 0.03875\text{mol} = 0.0775\text{mol of atoms}$$

Then number of Br atoms $= nN_A$

$$= 0.0775 \times 6.02 \times 10^{23} \text{ atoms} = 4.6655 \times 10^{22}\text{atoms}$$

Example 15

Calculate the number of calcium atoms in 10g of $Ca_3(PO_4)_2$.

Solution

Given mass of $Ca_3(PO_4)_2$ is 10g

Molar mass of $Ca_3(PO_4)_2$

$$= \big((3 \times 40) + (31 \times 2) + (16 \times 8)\big) \text{ g /mol} = 310\text{g/mol}$$

Mass of calcium (3mol of it) in 310g (1mol) of $Ca_3(PO_4)_2$

$$= 3 \times 40\text{g} = 120\text{g}$$

If we let mass of calcium in 10g of the compound to be **x** in grams, the mass of calcium in the given mass can simply be found as follows:

	Mass of calcium	Mass of $Ca_3(PO_4)_2$
In one mole of the compound	120g	310g
In 10g of the compound	**x**g	10g

From which; $310\mathbf{x} = 120 \times 10$; $\mathbf{x} =$ ` $\frac{1200}{310} = 3.87\text{g}$

Thus mass of calcium in 10g of the compound = 3.87g

Using $n = \frac{m}{M_r}$;

Number of moles of calcium atoms = $\frac{3.87 \text{ g}}{40 \text{ g/ mol}} = 0.09675\text{mol}$

Then from $N = n\,N_A$;

Number of moles calcium atoms

$$= 0.09675 \times 6.02 \times 10^{23} \text{ atoms} = 5.82 \times 10^{22} \text{ atoms}$$

Hence the number of calcium atoms in 10g of the compound is 5.82×10^{22} atoms

Alternative solution

Using $n = \frac{m}{M_r}$;

Number of moles of $Ca_3(PO_4)_2 = \frac{10g}{310 \text{ g / mol}} = 0.03226\text{mol}$

But total number moles of atoms in one mole of $Ca_3(PO_4)_2$

$$= \big(3 + (1 \times 2) + (4 \times 2)\big) \text{ moles} = 13 \text{ moles}$$

Then

$1 \text{ mol of } Ca_3(PO_4)_2 = 13 \text{ moles of atoms}$

$0.03226 \text{ mol of } Ca_3(PO_4)_3 = \mathbf{x} \text{ moles of atoms}$

And by cross multiplication;

$\mathbf{x} = 0.03226 \times 13 = 0.41938$

Thus number of moles of $Ca_3(PO_4)_2$ in 10g of it $= 0.41938\text{mol}$

Also $13\text{mol of atoms } Ca_3(PO_4)_3 \equiv 3\text{mol of Ca atoms}$

Then $0.41938 \text{ mol of atoms of } Ca_3(PO_4)_3 \equiv \text{xmol of Ca atoms}$

Again by cross multiplication;

$13\mathbf{x} = 0.41938 \times 3; \mathbf{x} = \frac{1.25814}{13} = 0.09678\text{mol}$

Thus number of moles of Ca atoms in 10g of the compound is 0.09678 mol

Finally using $N = n N_A$;

Number of Ca atoms

$=0.09678 \times 6.02 \times 10^{23}$ atoms $= 5.826 \times 10^{22}$atoms

Example 16

What mass of $FeSO_4.7H_2O$ do you need to dose anemic cat with 50mg of iron?

Solution

Given mass of iron = 50mg $= 50 \times 10^{-3}$g

Number moles of iron, $n = \frac{\text{Mass of iron}}{\text{Molar mass}}$

$$= \frac{50 \times 10^{-3}\text{g}}{56\text{ g /mol}} = 8.93 \times 10^{-4}\text{ mol}$$

1mole of $FeSO_4 \equiv$ 1mole of Fe

Thus number of moles of $FeSO_4$ that contains8.93×10^{-4} molof Fe will be also 8.93×10^{-4}mol

From $n = \frac{m}{M_r}$; $m = n \times M_r$;

Where M_r of $FeSO_4.7H_2O$

$= \big(56 + 32 + (16 \times 4)\big) + 7\big((1 \times 2) + 16\big)$= 278g / mol

$m = 8.93 \times 10^{-4}$ mol $\times 278$g/mol $= 0.25$ g

Mass of $FeSO_4.7H_2O$ needed is 0.25g

Alternative solution (By using the concept of percentage composition).

Mass percentage of Fe = $\left(\frac{\text{Mass of Fe in 1mol of } FeSO_4.7H_2O}{\text{molar mass of } FeSO_4.7H_2O}\right) \times 100\%$

But molar mass of $FeSO_4.7H_2O$

$$= \big(56 + 32 + (16 \times 4) + (7 \times 18)\big)\text{ g/mo} = 278\text{ g/mol}$$

And mass of Fe in 1mol of $FeSO_4.7H_2O = 56$ g

So % Fe $= \frac{56}{278} \times 100\% = 20.14\%$

With given mass of iron which is $50\text{mg} = 50 \times 10^{-3}\text{g}$

$$20.14 = \left(\frac{50 \times 10^{-3}}{\text{Mass of } FeSO_4 .7H_2O}\right) \times 100\%$$

From which;

$$\text{Mass of } FeSO_4 . 7H_2O = \frac{50 \times 10^{-3} \times 100}{20.14}\text{g} = 0.25\text{g}$$

Hence mass of $FeSO_4 . 7H_2O$ is 0.25g

Example 17

The density of O_2 at 0°C and 1atm is 1.429g/L. Calculate the molar volume of the gas.

Solution

Given density is 1.429g/L at 0°C and 1atm;

This implies that 1.429g of O_2 occupies the volume of 1L at 0°C and 1atm.

Thus, mass, m = 1.429g and Volume, V = 1L

From $n = \frac{m}{M_r}$ and $n = \frac{V}{GMV}$

Whence $n = \frac{m}{M_r} = \frac{V}{GMV}$

From which (by making GMV the subject)

$GMV = \frac{V \times M_r}{m}$; where M_r of $O_2 = 32$ g/mol

Thus $GMV = \frac{1L \times 32g/mol}{1.429g} = 22.4L/mol$

Hence the molar volume of the oxygen gas is 22.4L

Example 18

If the mole were defined to be 3×10^{24} instead of Avogadro's number; what would be the mass of one mole of argon atoms? Atomic weight of argon on conventional scale is 40.

Solution

Numerical value of atomic weight of Ar = Numerical value of molar mass

Mass of Ar atoms = 40g/mol (under conventional scale)

The mass of one atom of Argon = $\frac{40\text{gmol}^{-1}}{6.02\times10^{23}\text{ atom mol}^{-1}}$

$= 6.6445 \times 10^{-23}$ g/atom

Molar mass of atoms = mass of one atom × number of atoms in one mole

But the new number of atoms in one mole = 3×10^{24}atom/mol

Molar mass of Ar atoms

= 6.6445×10^{-23} g/atom $\times 3 \times 10^{24}$atom/mol= 199g/mol

Hence the new atomic weight of Ar is 199

Example 19

P and Q are two element which form P_2Q_3 and PQ_2 molecules. If 0.15mol of P_2Q_3 and PQ_2weighs 15.9g and 9.3g respectively. What are atomic mass of P and Q

Solution

Let x and y be atomic mass of P and Q respectively;

Finding molar mass of $P_2Q_3$by using $n = \frac{m}{M_r}$

From which $M_r = \frac{m}{n} = \frac{15.9g}{0.15g} = 106\text{g/mol}$

$M_r(P_2Q_3) = 106$

$2x + 3y = 106$…………………(i)

Also; from $M_r = \frac{m}{n}$

$$M_r(PQ_2) = \frac{9.3\text{g}}{0.15\text{ mol}} = 62\text{ g/mol}$$

Thus $x + 2y = 62$ ……………………..(ii)

Now we have the two equations:

$$2x + 3y = 106 \text{ and } x + 2y = 62$$

Solving the two equation simultaneously gives;

$x = 26$ and $y = 18$

Hence atomic mass of P is 26 and Q is 18

PRACTICE EXERCISE 4

Question 1

Calculate the number of moles in the following:

(a) 85.6g of CaO
(b) 0.547mg of $CuSO_4$
(c) 6.48g of $KMnO_4$
(d) 12.8g of NH_3

Question 2

How many silver atom are present in a ring of pure silver which has mass of 3g.

Question 3

How many atoms of copper are in a 33kg of pure copper statue?

(Molar mass of Cu = 64g/mol; molar mass Mn = 55g/mol)

Question 4

Calculate the number of moles in

(a) 3.25×10^{23} molecues of CO_2
(b) 7.5×10^{24} formula units of Ag_2CO_3

Question 5

How many atoms are contained in each of the following:

(a) 60.5g of $AlCl_3$
(b) 84.6 mL of HCl (g) at STP

Question 6

What is the mass of:

(i) Avogadro's number of carbon atoms?
(ii) 1.77×10^{30} molecules of CO_2

Question 7

Calculate the following:

(i) Mass (in grams) of 0.5mol of K_2SO_4
(ii) Number of moles present in 4.68g of aluminium sulphate.
(iii) Mass of 10^{23} molecules of water.

Question 8

In 2.5 mol of Ba $(NO_3)_2$;

(a) How many moles of barium ions present?
(b) How many moles of nitrate ions present?
(c) How many moles oxygen atom present?

Chapter 5

METHODS OF REPRESENTING CONCENTRATION

Concentration of a solution (also known as **strength of solution**) is the amount of the solute dissolved in a given amount of solution or solvent. Generally, concentration is the amount of a particular constituent of a mixture present in a given amount of the mixture.

When the solute is present in a very small quantity as compared with the solvent, the system is called **dilute solution**. The opposite of this is **concentrated solution.**

Concentration of solute in the solution can be expressed in various ways which will be discussed in this chapter.

MASS CONCENTRATION

This is the mass of the solute dissolved in a unit volume of solution.

That is: **Mass concentration** = $\frac{\textbf{Mass of solute}}{\textbf{Volume of solution}}$

Mass concentration is commonly given in g/dm^3,

Other units are g/cm^3, g/ml etc.

Relationship between mass concentrations of a solution components and density of the solution

By definition: Density of solution = $\frac{\text{Mass of solution}}{\text{Volume of solution}}$

But mass of solution = mass of solute + mass of solvent

Then: Density of solution $= \frac{\text{mass of solute+mass of solvent}}{\text{volume of solution}}$

$$= \frac{\text{Mass of solute}}{\text{volume of solution}} + \frac{\text{Mass of solvent}}{\text{Volume of solution}}$$

But $\frac{\text{mass of solute}}{\text{volume of solution}}$ = mass concentration of solute

And $\frac{\text{mass of solvent}}{\text{volume of solution}}$ = mass concentration of solvent

Hence:

Density of solution = mass concentration of solute + mass concentration of solvent

MOLAR CONCENTRATION (MOLARITY)

This is number of moles of solute dissolved in a litre of solution.

Thus $\textbf{Molarity (M)} = \frac{\textbf{number of moles of solute}}{\textbf{volume of solution in litres (or dm}^3\textbf{)}}$

Its unit is mol/L (or mol/dm^3) or M where $1\text{M} = 1\text{mol/dm}^3$

The reader should understand that: A molar solution is commonly used to mean that: One mole of solute is dissolved in a litre of solution.

Remember: 1dm^3 = 1Litre = 1000cm^3

Relationship between molarity and mass concentration

From Molarity $= \frac{\text{number of moles of solute}}{\text{volumeof solution in dm}^3}$

But number of moles of solute $= \frac{\text{mass of solute}}{\text{molar mass of solute}}$

Then: Molarity $= \frac{\text{mass of solute}}{\text{volume of solution in dm}^3 \times \text{molar mass of solute}}$

$\frac{\text{mass of solute}}{\text{volume of solution in dm}^3}$ = Mass of concentration in g/dm^3

Hence: $\textbf{Molarity} = \frac{\textbf{mass concentration in g/dm}^3}{\textbf{molar mass of solute in g/mol}}$

Or mass concentration in g/dm^3 = molarity × molar mass of solute in g/mol.

NORMALITY

This is the number of equivalents of solute dissolved in one litre of the solution

Thus: **Normality** $= \frac{\textbf{number of euivalents of solute}}{\textbf{volume of solution in litres (or dm}^3\textbf{)}}$

Its unit is N

Relationship between normality and molarity

From Normality $= \frac{\text{number of equivalents of solute}}{\text{volume of solution dm}^3}$

But number of equivalents of solute $= \frac{\text{mass of solute}}{\text{Equivalent weight of solute}}$

Then:

Normality $= \frac{\text{mass of solute}}{\text{volume of solution in dm}^3 \times \text{equivalent weight of solute}}$

But $\frac{\text{mass of solute}}{\text{volume of solution in dm}^3} =$ mass concentration of solute in g/dm^3

Thus: Normality $= \frac{\text{mass concentration in g/dm}^3}{\text{equivalent weight of solute}}$

Definition of equivalent weight

This is the mass of a compound required to give (or react with) one mole (Avogadro's number) of fundamental entity in a reaction we might consider.

For acid-base reaction:

Equivalent weight $= \frac{\text{molar mass}}{\text{acidity or basicity}}$

Then: Normality $= \frac{\text{mass concentration in g/dm}^3 \times \text{Basicity or Acidity}}{\text{molar mass of solute}}$

But $\frac{\text{mass concentration in g/dm}^3}{\text{molar mass of solute}}$ = molarity

Hence for acid – base reaction:

Normality = Molarity × Basicity (or Acidity)

The term **'A normal solution'** is used to express the solution with normality of 1N.

MASS PERCENTAGE $\left(\%\left(\frac{\mathbf{m}}{\mathbf{m}} \textbf{ or } \frac{\mathbf{w}}{\mathbf{w}}\right)\right)$

This is the mass of solute in grams dissolved in 100g of solution.

Thus $\%\left(\frac{\mathbf{m}}{\mathbf{m}}\right) = \frac{\textbf{mass of solute in g}}{\textbf{mass of solution in g}} \times \mathbf{100}\%$

The **mass percentage** (also termed as **weight percentage**) is the most common method of representing concentration as percentage, so whenever we talk (or you asked to find) just percentage without any extra information, the implication will always be mass percentage.

Relationship between mass percentage and mass concentration

From $\%\left(\frac{m}{m}\right) = \frac{\text{mass of solute in g}}{\text{mass of solution in g}}$

Where $\%\left(\frac{m}{m}\right)$ has been converted to fraction by dividing the percentage by 100 so that there is no need of multiplying the right hand side of the equation by 100 as shown in the original formula.

But mass of solution = Density of solution × volume of solution

(From Density = $\frac{\text{mass}}{\text{volume}}$; mass = Density × volume).

Then:

$\%\left(\frac{m}{m}\right) = \frac{\text{mass of solute in g}}{\text{volume of solution} \times \text{Density of solution}}$

But $\frac{\text{mass of solute in g}}{\text{volume of solution}}$ = mass concentration

Then $\%\frac{m}{m} = \frac{\text{mass concentration}}{\text{Density of solution}}$

Hence: **Mass concentration = Density of solution** $\times$**%** $\left(\frac{\mathbf{m}}{\mathbf{m}}\right)$

Where mass percentage, % $\left(\frac{m}{m}\right)$ must be converted to fraction by dividing the percentage by 100.

DILUTION PRINCIPLE

Dilution principle may be stated as:

Concentration of the solution containing fixed mass (or number of moles) of the solute varies inversely to its volume.

That is $M \propto \frac{I}{V}$ or MV = constant

From which $M_1V_1 = M_2V_2 = M_3V_3 = \ldots\ldots\ldots\ldots = M_nV_n$

To have better understanding this, suppose we have the following:

M_c = Molar concentration of concentrated solution.

M_d = Molar concentration of dilute solution.

V_c = Volume of the concentrated solution used to prepare the diluted solution of concentration, M_d

V_d = Volume of dilute solution obtaining after mixing pure water and the concentrated solution $\left(V_d = V_c + V_{H_2O}\right)$

Also using $M = \frac{n}{V}$ or $n = MV$

Number of moles of the solute in the concentrated solution = M_cV_c

Number of moles of the solute in the diluted solution = M_dV_d

Since the dilution is done by adding pure water (which has no solute in it), number of moles of solute in the concentrated solution and diluted solution remains constant and hence;

$$\mathbf{M_cV_c = M_dV_d}$$

So in other words, dilution principle may be stated as: *If the concentrated solution is diluted by adding more solvent, the number of moles of solute remains constant.*

This is true even if other methods of representing concentration (apart from molarity) are used.

So generally;

$$\mathbf{Q_cC_c = Q_dC_d}$$

Where:

$\mathbf{C_c}$and$\mathbf{C_d}$are concentration (mass concentration, molarity, normality or mass percentage) of concentrated and diluted solution respectively.

$\mathbf{Q_c}$and $\mathbf{Q_d}$are quantity of concentrated and diluted solution respectively (may be volume if mass concentration, molarity or normality is used or may be mass if mass percentage is used).

Example 1

What mass of solute is present in each aqueous solution?

(a) 40cm^3 of 6mol/dm^3 sulphuric acid solution
(b) 4.24dm^3 of 0.775mol/L ammonium nitrate solution

Solution

(a) Volume of H_2SO_4 = $40\text{cm}^3 = 40 \times 10^{-3}\text{dm}^3$
Molarity of H_2SO_4 = 6mol/dm^3
Mass of solute =?

Using $\text{n} = \text{MV}$and also $\text{n} = \frac{\text{m}}{\text{M}_\text{r}}$;

Then by equating; $\frac{\text{m}}{\text{M}_\text{r}} = \text{MV}$;

From which $\text{m} = \text{MVM}_\text{r}$;

Where $\text{M}_\text{r}(\text{H}_2\text{SO}_4) = (1 \times 2) + 32 + (16 \times 4) = 98\text{g/mol}$

Mass of H_2SO_4 $= 40 \times 10^{-3}\text{dm}^3 \times 6\text{moldm}^{-3} \times 98\text{g} = 23.52\text{g}$

(b) Volume of ammonium nitrate = $4.24\mathrm{dm}^3$

Molarity of ammonium nitrate = $0.775\mathrm{mol/L}$

Using $\mathrm{n} = \mathrm{MV}$ and also $\mathrm{n} = \frac{\mathrm{m}}{\mathrm{M_r}}$;

Then by equating; $\frac{\mathrm{m}}{\mathrm{M_r}} = \mathrm{MV}$;

From which $\mathrm{m} = \mathrm{MVM_r}$;

Where$\mathrm{M_r}(\mathrm{NH_4NO_3}) = (14 \times 2) + (4 \times 1) + (16 \times 3) = 80\mathrm{g/mol}$

Mass of $\mathrm{NH_4NO_3} = 4.24\mathrm{dm}^3 \times 0.775\mathrm{moldm}^{-3} \times 80\mathrm{g} = 262.88\mathrm{g}$

Example 2

What volume of $0.250\ \mathrm{mol/L}$ solution can be made by using 14g of sodium hydroxide?

Solution

Molarity of sodium hydroxide, $\mathrm{M} = 0.25\mathrm{mol/L}$

Mass of sodium hydroxide $\mathrm{M} = 14\mathrm{g}$

Volume sodium hydroxide, V = ?

Molar mass of sodium hydroxide = $23 + 16 + 1 = 40\mathrm{g/mol}$

From$\mathrm{M} = \frac{\mathrm{n}}{\mathrm{V}}$; $\mathrm{M} = \frac{\mathrm{m}}{\mathrm{M_r} \times \mathrm{V}}$

From which(by making V the subject)

$$\mathrm{V} = \frac{\mathrm{m}}{\mathrm{M} \times \mathrm{M_r}} = \frac{14\mathrm{g}}{0.25\mathrm{mol/L} \times 40\mathrm{g/mol}} = 1.4\mathrm{L}$$

Volume of sodium hydroxide = 1.4L

Alternative solution

By equating the two equations $\mathrm{n} = \mathrm{MV}$ and $\mathrm{n} = \frac{\mathrm{m}}{\mathrm{M_r}}$; $\mathrm{MV} = \frac{\mathrm{m}}{\mathrm{M_r}}$

Then substituting $0.25\mathrm{molL}^{-1} \times \mathrm{V} = \frac{14\mathrm{g}}{40\mathrm{gmol}^{-1}}$;

From which $V = \frac{14g}{0.25mol/L \times 40g/mol} = 1.4L$

Example 3

A 100mL bottle of skin lotion contains a number of solutes. One of these solutes is zinc oxide. The concentration of zinc oxide in the skin lotion is $0.915mol/L$. What mass of zinc oxide is present in the bottle?

Solution

Volume of solution, $V = 100mL = 100 \times 10^{-3}L = 0.1L$

Concentration of zinc oxide, $M = 0.915mol/L$

Mass of zinc oxide = ?

Molar mass of Zinc oxide = $ZnO = 65 + 16 = 81g/mol$

From $M = \frac{n}{V}; n = MV$

Also $n = \frac{m}{M_r}$

Thus (by equating); $MV = \frac{m}{M_r}$

From which;

$m = M \times V \times M_r = 0.915mol/L \times 81g/mol \times 0.1L = 7.41g$

Hence mass of zinc oxide is 7.41g

Example 4

Determine mass of $Na_2SO_4 .10H_2O$ required for preparing 250mL of salt solution whose molarity is 0.45M.

Solution

Molarity of $Na_2SO_4.10H_2O$ solution $= 0.45M$

Volume of $Na_2SO_4.10H_2O$ solution $= 250mL = 250 \times 10^{-3}L$

Using $M = \frac{n}{V}$; But $n = \frac{m}{M_r}$;

Thus $M = \frac{m}{M_r \times V}$

From which (by making m the subject)

$m = M \times M_r \times V$; Where $M_r(Na_2SO_4.10H_2O) = 322g/mol$

Thus mass of $Na_2SO_4.10H_2O$

$= 0.45mol/L \times 322g/mol \times 250 \times 10^{-3}L = 3.623 \times 10g = 36.23g$

Hence the required mass is $36.23g$

Example 5

Find volume of 0.2N solution containing 2.5meq of solute.

Solution

Number of equivalents = $2.5meq = 2.5 \times 10^{-3}eq$ (meq stands for milliequivalent whereby $1meq = 10^{-3}eq$)

From normality $= \frac{\text{Number of equivalents}}{\text{Volume of solution}}$;

$$\text{Volume of solution} = \frac{\text{Number of equivalents}}{\text{Normality}}$$

$$\text{Volume of solution} = \frac{2.5 \times 10^{-3}eq}{0.2} = 0.0125dm^3$$

Hence the volume of solution $0.0125dm^3$

Example 6

$0.65M\ BaCl_2$ solution is prepared by dissolving pure solid $BaCl_2.2H_2O$ in water. Determine the mass of hydrated salt dissolved per millilitre of solution and mass of anhydrous $BaCl_2$ present per millilitre of solution.

Solution

Molarity of solution, M $= 0.65\text{M} = 0.65\text{mol/dm}^3$

Volume, V = 1mL= 10^{-3}L

Using $\text{M} = \frac{\text{n}}{\text{V}}$; But $\text{n} = \frac{\text{m}}{\text{M}_\text{r}}$;

Thus $\text{M} = \frac{\text{m}}{\text{M}_\text{r} \times \text{V}}$

From which (by making m the subject)

$\text{m} = \text{M} \times \text{M}_\text{r} \times \text{V}$; where $\text{M}_\text{r}(\text{BaCl}_2.2\text{H}_2\text{O}) = 244\text{g/mol}$

Mass of pure $\text{BaCl}_2.2\text{H}_2\text{O}$

$= 0.65\text{mol /L} \times 244\text{g/mol} \times 10^{-3}\text{L} = 0.1586\text{g}$

Hence the mass of hydrated salt is 0.1586g

Molar mass of anhydrous $\text{BaCl}_2 = 208\text{g/mol}$

Again using $\text{m} = \text{M} \times \text{M}_\text{r} \times \text{V}$ and the fact that;

$$\text{Molarity of hydrated salt} = \text{Molarity of anhydrous salt}$$

Mass of $\text{BaCl}_2 = 0.65\text{mol/L} \times 208\text{g/mol} \times 10^{-3}\text{L} = 0.135\text{g}$

Hence the mass of anhydrous salt is 0.1352g

Example 7

Density of sulphuric acid solution is 1.2g/mL and it is 40% H_2SO_4. Determine molar concentration of this solution.

Solution

From mass concentration = Density × % $(^\text{m}/_\text{m})$

Given density $= 1.2\text{g/mL} = \frac{1.2\text{g}}{10^{-3}\text{L}} = 1200\text{g/L}$

Mass concentration = $1200\text{g/L} \times \frac{40}{100} = 480\text{g/L}$

But molarity = $\frac{\text{Mass conentration}}{\text{Molar mass}} = \frac{480\text{g/L}}{M_r(H_2SO_4)}$

Where $M_r(H_2SO_4) = \left((1 \times 2) + 32 + (16 \times 4)\right) = 98\text{g/mol}$

Thus molarity = $\frac{480\text{g/L}}{98\text{g/mol}} = 4.9\text{mol/L}$

Hence molarity of the solution is 4.9M

Example 8

An aqueous solution is prepared by dissolving pure crystals Mohr's salt, $FeSO_4.(NH_4)_2SO_4.6H_2O$ in water. Density of the solution is 1.2g/mL and the solution contain 30%. $FeSO_4.(NH_4)_2SO_4$ by weight. Determine molarity of this solution and mass of the salt dissolved if the volume of solution is 400mL.

Solution

From molarity = $\frac{\text{Mass concetration}}{\text{Molar mass}}$

But mass concentration = $\text{density of solution} \times \%(\text{m/m})$

Molarity = $\frac{\text{Density of solution} \times \%(\text{m/m})}{\text{Molar mass}}$

Where molar mass of $FeSO_4.(NH_4)_2SO_4.6H_2O$ = 392g/mol

Density of $FeSO_4.(NH_4)_2SO_4.6H_2O$ is $1.2\text{g/mL} = 1200\text{g/L}$

Mass to mass percentage $\%(\text{m/m}) = 30\%$

Therefore molarity = $\frac{1200\text{g/L} \times {}^{30}/_{100}}{392\text{g/mol}} = 0.918\text{mol/L}$

Hence molarity of the solution is 0.918mol/L

Also using $M = \frac{n}{V}$; But$n = \frac{m}{M_r}$;

Thus $M = \frac{m}{M_r \times V}$

From which (by making m the subject)

$m = M \times M_r \times V$; where volume of solution = $400mL = 0.4L$

Thus mass of the salt = $0.918 \times 0.4 \times 392g = 143.9g$

Hence the mass of the salt dissolved is $143.9g$

Example 9

How many grams wet NaOH containing 10% water are required to prepare 1 litre of 0.1N solution?

Solution

$$\text{Molarity of NaOH} = \frac{\text{Normality of NaOH}}{\text{Acidity of NaOH}}$$

But acidity of NaOH is 1

Thus molarity of NaOH = $\frac{0.1}{1}M = 0.1M$

Using $n = MV$;

Number of moles of NaOH in 1L

= $0.1 \text{ mol dm}^3 \times 1dm^3 = 0.1 \text{ mol}$

Using $m = n\,M_r$; (Where M_r of NaOH $(23 + 16 + 1)g/mol = 40g/mol$)

Mass of pure NaOH $= 0.1mol \times 40g/mol = 4g$

$$\text{But \% pure NaOH} = \frac{\text{mass of pure NaOH}}{\text{mass of wet (impure)NaOH } (m_w)} \times 100\%$$

And;

$$\%\text{pure NaOH} = 100\% - \%H_2O(\%\text{pure NaOH} + \%H_2O = 100\%)$$
$$= (100 - 10)\% = 90\%$$

Substituting $90 = \frac{4g}{m_w} \times 100$; $m_w = \frac{4 \times 100}{90} = 4.44g$

Hence mass of wet NaOH required is 4.44g

Example 10

Calculate mass percentage of potassium chloride in a solution prepared by mixing 8g of KCl with 42g of water.

Solution

Mass percentage $\% \left({}^{m}/{}_{m}\right) = \frac{\text{Mass of solute}}{\text{Mass of solution}} \times 100\%$

Mass of solute = 8g

Mass of solution = 42 + 8g = 50g

$\% \left({}^{m}/{}_{m}\right) = \frac{8g}{50g} \times 100\% = 0.16 \times 100\% = 16\%$

Therefore the mass percentage is $16\% \left({}^{m}/{}_{m}\right)$

Example 11

What is molarity of a 0.35L solution that contains 46g of potassium chloride?

Solution

Mass of potassium chloride = 46g

Volume of potassium chloride = 0.35L

Molarity = $\frac{\text{number of moles}}{\text{Volume in litres}} = \frac{\text{Mass}}{\text{Molar mass} \times \text{Volume}}$

Molar mass, KCl $= 39 + 35.5 = 74.5\text{g/mol}$

Thus molarity = $\frac{46g}{74.5\text{g/mol} \times 0.35L} = 1.76\text{mol/L}$

Molarity of KCl is $1.76\ \text{mol/L}$

Example 12

How many grams of sodium chloride, are needed to prepare 225g of a $10\% \left({}^{m}/{}_{m}\right)$ NaCl solution?

Solution

$$\%(^{m}/_{m}) = \frac{\text{Mass of solute}}{\text{Mass of solution}} \times 100\%$$

$$\frac{\text{Mass of solute}}{\text{Mass of solution}} \times 100\% = 10\%$$

$$\frac{\text{Mass of solute}}{\text{Mass of solution}} = \frac{10}{100}$$

$$\text{Mass of solute} = \frac{10}{100} \times \text{Mass of solution}$$

$$= \frac{10}{100} \times 225\text{g} = 22.5\text{g}$$

Mass of NaCl is 22.5g

Example 13

How many litres of 2M NaCl solution are needed to provide 67.3g of NaCl ?

Solution

$$\text{From molarity} = \frac{\text{number of moles}}{\text{Volume of solution in dm}^3};$$

$$\text{Molarity} = \frac{\text{Mass}}{\text{Molar mass} \times \text{Volume}}$$

$$\text{From which volume} = \frac{\text{Mass}}{\text{Molarity} \times \text{Molar mass}}$$

$$= \frac{67.3\text{g}}{2\text{mol/dm}^3 \times 58.5\text{g/mol}} = 0.575\text{dm}^3$$

The volume is 0.575dm^3

Alternative solution

$$\text{From } M = \frac{n}{V};\ n = MV;$$

$$\text{Also } n = \frac{m}{M_r}$$

Thus (by equating); $MV = \frac{m}{M_r}$

Then substituting $2moldm^{-3} \times V = \frac{67.3g}{58.5gmol^{-1}}$;

From which $V = \frac{67.3g}{2mol/dm^3 \times 58.5g/mol} = 0.575dm^3$

Example 14

What is the molarity of solution prepared by diluting $10cm^3$ of $2.7M$ NaOH to $60cm^3$

Solution

Volume of concentrated solution, $V_c = 10cm^3$

Molarity of concentrated solution, $M_c = 2.7M$

Volume of diluted solution, $V_d = 60cm^3$

Molarity of diluted solution. $M_d = ?$

From $M_cV_c = M_dV_d$;

$$M_d = \frac{M_cV_c}{V_d} = \frac{2.7M \times 10cm^3}{60cm^3} = 0.45M$$

Molarity of the solution is$0.45M$

Example 15

A solution is prepared by adding $600cm^3$ of distilled water to $100cm^3$ of $0.15mol/dm^3$ ammonium nitrate (NH_4NO_3). Calculate the molar concentration of solution.

Solution

Molarity concentrated solution,$M_c = 0.15mol/dm^3$

Volume of concentrated solution, $V_c = 100cm^3$

Volume of dilute solution = V_c + volume of distilled water

$$= 100cm^3 + 600cm^3 = 700cm^3$$

From dilution formula; $M_d V_d = M_c V_c$

$$M_d = \frac{M_c V_c}{V_d} = \frac{100 \times 0.15 mol/dm^3}{700} = 0.021 mol/dm^3$$

Hence the molar concentration is $0.021\ mol/dm^3$

Example 16

80g of $60\% HNO_3 (m/m)$ is diluted to$20\% HNO_3 (m/m)$. Calculate mass of water added.

Solution

Using dilution formula;

$$Q_c C_c = Q_d C_d$$

Where:

$Q_c = 80g, C_c = 60\%(m/m), \qquad C_d = 20\%(m/m)$

Substituting $80g \times 60\%(m/m) = Q_d \times 20\%(m/m)$

From which $Q_d = \frac{80g \times 60\%(m/m)}{20\%(m/m)} = 240g$

Thus mass of diluted solution $= 240g$

But; mass of diluted solution

$= \text{mass of concentrated solution} + \text{mass of water}$

That is $240 = 80 + \text{mass of water}$;

From which mass of water $= 240 - 80 = 160g$

Hence mass of water added is $160g$.

PRACTICE EXRECISE 5

Question 1

A solution is prepared with 15g of sodium carbonate and 235g of water. What is the mass percentage of sodium carbonate in the solution?

Question 2

Calculate the mass and molar concentration and mass percentage of each of the following:

(a) 63g of HNO_3 is dissolved in enough water to make 100 litres of solution.
(b) 45g of glucose, $C_6H_{12}O_6$ is dissolved in enough water to make 05L of solution.
(c) 116g of potassium fluoride is dissolved in enough water to make 4L of solution.

Question 3

You had 60g of 20% coal tar solution which was diluted to produce 100g. What is the strength of the final product?

Question 4

Calculate the mass / mass percent of solute for each solution.

(a) 17g of sulphuric acid and dissolved in 65g of the solution.
(b) 18.37g of sodium chloride dissolved in 92.2g of water.

Question 5

Steel is an alloy of iron and contain about 7% carbon. It is also containing small amount of other materials, such as manganese and phosphorous. What mass of carbon is needed to make a 5kg sample of steel?

Question 6

A student dissolved 30.46g of silver nitrate in water to make 500cm^3 of solution. What is the molar concentration of the solution?

Question 7

Formalin is an aqueous solution of formaldehyde HCHO, used to preserve biological specimen. What mass of formaldehyde is needed to prepare 1.5L of formalin with a concentration of 10mol/L?

Question 8

Suppose that you are given a solution of 1.25mol/L sodium chloride in water. What volume must you dilute to prepare the following solution?

(a) 20mL of 0.8M NaCl
(b) 250mL of 0.3M NaCl

Question 9

How many cm^3 of 18M Sulphuric acid are required to prepare 300cm^3 of 1M H_2SO_4?

Question 10

What volume of water should be added to 100 cm^3 sulphuric acid having specific gravity 1.1 and 80% strength to give 1N solution?

Chapter 6

EMPIRICAL AND MOLECULAR FORMULA

Empirical formula *is the chemical formula which gives the simplest whole number ratio of number of atoms present in a compound.*

Molecular formula *is the chemical formula which gives exact number of each atom present in a molecule of a compound.*

So the empirical formula gives the whole number ratio of number of atoms of different elements in the compound while molecular formula is the whole number multiple of empirical formula.

That is if A_xB_Y is the empirical formula;

Then the molecular formula will be $(A_xB_Y)_n = A_{xn}B_{Yn}$ where n is the positive whole number. That is n may be 1, 2, 3, 4 etc.

(If n = 1, the empirical formula become the same as the molecular formula).

To have better understanding of the two terms consider the compound, butene whose molecule consist of four carbon atoms and eight hydrogen atoms.

- Thence the molecular formula of butene will be C_4H_8 (it gives exact number of each atom present in one molecule butene).
- Its empirical formula is then found by determining the simplest ratio of 4 and 8 (number of carbon atoms and hydrogen atoms respectively) is done by dividing each number of atoms by 4 which is actually the greatest common factor (GCF) of 4 and 8.

That is $\frac{4\text{ carbon}}{8\text{ hydrogen}} \div \frac{4}{4} = \frac{1\text{ carbon}}{2\text{ hydrogen}}$

Hence the empirical formula of butene will be CH_2

Don't use the term molecular formula for ionic compound! Molecular formula is for covalent compounds only because only covalent compounds consist of molecules.

- Ionic compound consists of infinite number (too many) of formula units in the ionic lattice. So unlike covalent compounds which are commonly represented by using molecular formula, ionic compounds are always represented by using empirical formula. For example, sodium chloride is always represented as NaCl (one unit of sodium chloride consists of one sodium ion and one chloride ion).

Example 1

In each case below, the molecular formula for compound is given. Determine the empirical formula for each compound.

(a) C_3H_6(b) $C_6H_4Cl_4O_2$(c) Cl_2O_5

Answer

(a) The largest number that can divide each atom present in C_3H_6 is 3 (GCF of 3 and 6)

So dividing each number of atom present by 3;

$$\frac{3C}{3} = 1C; \frac{6H}{3} = 2H$$

Thus C_3H_6 (M.F) = $(CH_2)_3(E.F \times 3)$

Hence the empirical formula is CH_2

(b) The largest number that can divide each atom present in $C_6H_4Cl_4O_2$ is 2

Divide each atom present by 2;

$$\frac{6 \text{ carbons}}{2} = 3 \text{ carbons}$$

$$\frac{4 \text{ hydrogens}}{2} = 2 \text{ hydrogens}$$

$$\frac{2 \text{ Oxygen}}{2} = 1 \text{ oxygen}$$

So the compound will be $C_3H_2Cl_2O$

Alternative solution

Using the fact that molecular formula is whole number multiple of empirical formula;

$$C_6H_4Cl_4O_2 = (C_3H_2Cl_2O)_2$$

Hence the empirical formula is $C_3H_2Cl_2O$

(c) Cl_2O_5, there is no number greater than 1 (GCF of 2 and 5) that can divide each atom present in Cl_2O_5. Its empirical formula will be the same as molecular formula.

Therefore the empirical formula is Cl_2O_5

Example 2

Determine molecular formula of the following:

(a) A white powder which is analyzed and found to have the empirical formula of P_2O_5. The compound has a molar mass of 284g/mol.
(b) A compound with an empirical formula of NH_2 and a formula mass of 32amu.

Solution

(a) From the given empirical formula of P_2O_5.

Molecular formula will be $(P_2O_5)_n$ where n is the positive whole number.

So to get n: is from molar mass of compound which is 284g/mol

$$284 = M_r((P_2O_5)_n); \big((31 \times 2) + (16 \times 5)\big)n = 284$$

$$142n = 284; n = 2$$

So the molecular formula will be $(P_2O_5)_2 = P_4O_{10}$

(b) Let the molecular formula be $(NH_2)_n$

Then $14n + 2n = M_r = 32; 16n = 32, n = 2$

Hence the molecular formula isN_2H_4

Example 3

A compound used as an additive for gasoline to help prevent engine knocks shows the following percentages:

71.65% Cl, 24.27% C, 4.07% H. The molar mass of the compound is $99g/mol$. Determine the empirical formula and molecular formula for this compound.

Solution

Constituent atoms	**Cl**	**C**	**H**
Percentage composition of each	71.65	24.27	4.07
Mass of each in 100g of compound	71.65g	24.27g	4.07g
Number of mole of each; using $n = \frac{m}{M_r}$	$\frac{71.65g}{35.5gmol^{-1}}$ $= 2$ mol	$\frac{24.27g}{12gmol^{-1}}$ $=2.02$ mol	$\frac{4.07g}{1gmol^{-1}}$ $= 4.07$ mol
Dividing by smallest number to get simpler ratio	$\frac{2mol}{2mol}$ $= 1$	$\frac{2.02mol}{2mol}$ $= 1.01$ ≈ 1	$\frac{4.07\ mol}{2mol}$ $= 2.035$ ≈ 2

Hence the empirical formula is $ClCH_2 = CH_2Cl$

Let molecular formula be $(CH_2Cl)_n$

Molar mass of $(CH_2Cl)_n$ = Molar mass of compound

$$(12 + 2 + 35.5)n = 99$$

$$49.5n = 99;\ n = \frac{99}{49.5} = 2$$

Hence the molecular formula is $C_2H_4Cl_2$

Example 4

A compound consists of 40% carbon, 6.713% hydrogen and 53.28% oxygen on a mass basis; and has a molar mass of approximately 180g/mol. Determine the molecular formula and hence empirical formula.

Solution

In this question it is required to find first molecular formula before empirical formula.

So unlike the previous example which we used mass of each atom in 100g of compound, now we are going to find mass of each atom in 180g (molar mass of the compound) so as to find the molecular formula directly.

For carbon (40%): mass of carbon in 180g of the compound

$$= \frac{40}{100} \times 180g = 72g$$

For hydrogen (6.713%): mass of hydrogen in 180g of the compound $= \frac{6.713}{100} \times 180g = 12.08g$

For oxygen (53.28%): mass of oxygen in 180g of the compound

$$= \frac{53.28}{100} = 180g = 95.904g$$

Constituent atoms	**C**	**H**	**O**
Mass of each in 180g of compound	72g	12.08g	95.904g
Number of moles of each;$n = \frac{m}{M_r}$	$\frac{72g}{12g/mol}$ $= 6\ mol$	$\frac{12.08g}{1g/mol}$ $= 12.08mol \approx 12\ mol$	$\frac{95.904g}{16g/mol}$ $= 5.994mol \approx 6\ mol$

Thus in one mole of the compound there are 6mol of carbon atoms, 12mol of hydrogen atoms and 6mol of oxygen atoms. Therefore, in one molecule of the compound there are 6 atoms of carbon, 12 atoms of hydrogen and 6 atoms of oxygen.

Hence the molecular formula is $C_6H_{12}O_6$

The largest number that can divide all number of atoms in the obtained molecular formula is 6;

Thus $C_6H_{12}O_6 = (CH_2O)_6$

Hence the empirical formula is CH_2O

Example 5

A compound was found to contain 33% silicon and 67% of chlorine, by mass. If the vapour density of the compound is 85, calculate its molecular formula.

Solution

Constituent atom	Si	Cl
Percentage composition	33	67
Mass of each in 100g of the compound	33g	67g
Number of moles of each; $n = \frac{m}{M_r}$	$\frac{33g}{28gmol^{-1}}$ $= 1.1786mol$	$\frac{67}{35.5gmol^{-1}}$ $= 1.8873\ mol$
Dividing by smallest to gest simpler ratio	$\frac{1.1786\ mol}{1.1786\ mol} = 1$	$\frac{1.8873\ mol}{1.1786\ mol} = 1.6$
Multiplying by 2 to get simplest ratio	$1 \times 2 = 2$	$1.6 \times 2 = 3.2 \approx 3$

Empirical formula is Si_2Cl_3

Let the molecular formula be $(Si_2Cl_3)_n$

But $M_r = \text{Vapour density} \times 2 = 85 \times 2 = 170$

Thus$\big((28 \times 2) + (3 \times 35.5)\big)n = M_r = 170;$

$162.5n = 170;\ n = \frac{170}{162.5} = 1.046 \approx 1$

Molecular formula is $(Si_2Cl_3)_n = (Si_2Cl_3)_1 = Si_2Cl_3$

Hence the molecular formula is Si_2Cl_3

Alternative solution: We may find molecular formula directly without finding empirical formula as follows:

Constituent atom	**Si**	**Cl**
Percentage composition	33	67
Mass of each in 170g of the compound	$\frac{33}{100} \times 170\text{g}$ $= 56.1\text{g}$	$\frac{67}{100} \times 170\text{g}$ $= 113.9\text{g}$
Number of moles of each; $n = \frac{m}{M_r}$	$\frac{56.1\text{g}}{28\text{gmo}^{-1}} = 2$	$\frac{113.9\text{g}}{35.5\text{gmo}^{-1}} = 3$

Hence the molecular formula is Si_2Cl_3

Example 6

A compound was analysed and found to contain$13.5\text{g Ca}, 10.8\text{g O}$ and 0.675g H. What is empirical formula of the compound?

Solution

Finding the percentage composition of each in the compound:

Total mass = $13.5\text{g} + 10.8\text{g} + 0.675\text{g} = 24.975\text{g}$

$$\%\ \text{Ca} = \frac{13.5\text{g}}{24.975\text{g}} \times 100 = 54.05\%$$

$$\%\ \text{O} = \frac{10.8\text{g}}{24.975\text{g}} \times 100 = 43.24\%$$

$$\%\ \text{H} = \frac{0.675\text{g}}{24.975\text{g}} \times 100 = 2.7\%$$

Constituent atoms	**Ca**	**O**	**H**
Percentage composition	54.05	43.24	27
Mass of each in 100g of compound	54.05g	43.24g	2.7g
Number of moles of each; $n = \frac{m}{M_r}$	$\frac{54.05g}{40g/mol} = 1.3513\ mol$	$\frac{43.24g}{16g/mol} = 2.7025mol$	$\frac{2.7g}{1mol} = 2.7\ mol$
Dividing by the smallest number to get simpler ratio	$\frac{1.3513\ mol}{1.3513\ mol}$ 1	$\frac{2.7025\ mol}{1.3513\ mol}$ 2	$\frac{2.7\ mol}{1.3513\ mol}$ 2

The empirical formula isCaO_2H_2

Alternative solution

Constituent atom	**Ca**	**O**	**H**
Mass of each in the sample	13.5g	10.8g	0.675g
Number of mole of each ; $n = \frac{m}{M_r}$	$\frac{13.5g}{40g/mol} = 0.3375\ mol$	$\frac{10.8g}{16g/mol} = 0.675\ mol$	$\frac{0.675g}{1g/mol} = 0.675\ mol$
Dividing by the smallest number to get simpler ratio	$\frac{0.3375\ mol}{0.3375\ mol} = 1$	$\frac{0.675\ mol}{0.3375\ mol} = 2$	$\frac{0.675\ mol}{0.3375\ mol} = 2$

The empirical formula is CaO_2H_2 (Which is actually $Ca(OH)_2$)

Example 7

An organic compound has molar mass of 150g/mol and contains 72%carbon, 6.67% hydrogen and 21.33%oxygen. Find empirical formula and hence molecular formula of the compound.

Solution

Constituents atoms	C	H	O
Percentage composition of each	72	6.67	21.33
Mass of each in 100g of the compound	72g	6.67g	21.33g
Number of moles of each, Using $n = \frac{m}{M_r}$	$\frac{72g}{12gmol^{-1}} = 6mol$	$\frac{6.67g}{1gmol^{-1}} = 6.67mol$	$\frac{21.33g}{16gmol^{-1}} = 1.33mol$
Dividing by smallest number to get simpler ratio	$\frac{6mol}{1.33mol} = 4.5mol$	$\frac{6.67mol}{1.33mol} = 5mol$	$\frac{1.33mol}{1.33mol} = 1$
Multiplying by 2 throughout to get simplest ratio	9	10	2

Thus the empirical formula for the compound is $C_9H_{10}O_2$

Let Molecular formula be $(C_9H_{10}O_2)_n$

Then 108n+10n+32n= M_r; 150n=150 or n=1

Hence molecular formula of the compound is $C_9H_{10}O_2$

Example 8

When heated in the presence of oxygen, 12g of the magnesium form an oxide with mass of 20g. Find the empirical formula of the oxide.

Solution

The oxide will consist of magnesium and oxygen.

Thusmass of oxygen $=$ Mass of oxide $-$ Mass of Magnesium

$$= 20g - 12g = 8g$$

Therefore, in 20g of the sample of the oxide:

Mass of magnesium $=$ 12g

Mass of oxygen $=$ 8g

Constituent atom	**Mg**	**O**
Mass of each in 20g of the compound	12g	8g
Number of mole of each; $n = \frac{m}{M_r}$	$\frac{12g}{24g/mol}$ $= 0.5$ mol	$\frac{8g}{16\ g/mol}$ $= 0.5$ mol
Divide by the smallest number to get simpler ratio	$\frac{0.5\ mol}{0.5\ mol}$ $= 1$	$\frac{0.5\ mol}{0.5\ mol}$ $= 1$

Therefore, the empirical formula is MgO.

Example 9

A sample of gas occupies 2 litre at STP. The sample contains 2.143g of carbon and 0.358g hydrogen. Find the empirical and molecular formula of the gas.

Solution

Constituent atoms	C	H
Mass of each in the sample	2.143g	0.358g
Number of moles of each atom; $n = \frac{m}{M_r}$	$\frac{2.143g}{12gmol^{-1}}$ $= 0.1786\ mol$	$\frac{0.358g}{1gmol^{-1}}$ $= 0.358mol$
Dividing by the smallest number to get simpler ratio	$\frac{0.1786\ mol}{0.1786\ mol}$ $= 1$	$\frac{0.358\ mol}{0.1786\ mol}$ $= 2$

Hence empirical formula is CH_2

Using $n = \frac{V}{GMV} = \frac{m}{M_r}$; where $GMV = 22.4\ dm^3/mol$ (at STP) and $m = 2.143g + 0.358g = 2.501g$

Substituting $\frac{2\ dm^3}{22.4\ dm^3/mol} = \frac{2.501\ g}{M_r}$

From which $M_r = \frac{2.501\ g \times 22.4\ dm^3/mol}{2\ dm^3} = 28g/mol$

Let the molecular formula be $(CH_2)_n$

$(CH_2)_n = M_r = 28; \left(12 + (1 \times 2)\right)n = 28$

$14n = 28; n = 2$

Molecular formula is $(CH_2)_n = (CH_2)_2 = C_2H_4$

Therefore molecular formula is C_2H_4

Example 10

A 60g tetraethyl lead a gasoline additive is found to contain 38.43g lead, 17.83g carbon and 3.74g hydrogen.

(i) Determine percentage composition
(ii) Find the empirical formula of the compound

Solution

Percentage composition of each:

Using percentage of element $= \frac{\text{mass of element}}{\text{mass of the compound}} \times 100$

For lead: $\frac{38.43\text{g}}{60\text{g}} \times 100\% = 64.05\%$

For carbon: $\frac{17.83\text{g}}{60\text{g}} \times 100\% = 29.72\%$

For hydrogen: $\frac{3.74\text{g}}{60\text{g}} \times 100\% = 6.23\%$

Hence the compound is 64.05%lead, 29.72% carbon and 6.23% hydrogen.

Constituent atom	**Pb**	**C**	**H**
Percentage Composition	64.05	29.72	6.23
Mass of each in 100g of the compound	64.05g	29.72g	6.23g
Number of mole of each; $n = \frac{m}{M_r}$	$\frac{64.05\text{g}}{207\text{gmol}^{-1}}$ $= 0.309$ mol	$\frac{29.72\text{g}}{12\text{gmol}^{-1}}$ $= 2.477$ mol	$\frac{6.23}{1\text{gmol}^{-1}}$ $= 6.23$ mol
Dividing by smallest number to get simpler ratio	$\frac{0.309 \text{ mol}}{0.309 \text{ mol}}$ $= 1$	$\frac{2.477 \text{ mol}}{0.309 \text{ mol}}$ $= 8$	$\frac{6.23 \text{ mol}}{0.309 \text{ mol}}$ $= 20$

Hence the empirical formula is PbC_8H_{20}

Example 11

A compound contains C, H and O. A 5g sample contains 3.215g carbon, and 0.357g hydrogen. A gaseous sample of this gas has a mass of 5.25 times the mass of an identical volume of oxygen gas under the same conditions. What is the molecular formula of compound?

Mass of C + Mass of H + Mass of O = 5g

3.215g + 0.357g + Mass O = 5g

Mass of O = 5g – 3.572=1.428g

Solution

Constituent atoms	**C**	**H**	**O**
Mass of each in 5g of the compound	3.215g	0.357g	1.428g
Number of mole of each; $n = \frac{m}{M_r}$	$\frac{3.215g}{12gmol^{-1}} = 0.2679mol$	$\frac{0.357g}{1gmol^{-1}} = 0.357mol$	$\frac{1.428g}{16gmol^{-1}} = 0.089mol$
Dividing by smallest number	$\frac{0.2679mol}{0.089} = 3$	$\frac{0.357mol}{0.089mol} = 4$	$\frac{0.089mol}{0.089mol} = 1$

Empirical formula is C_3H_4O

Under the same condition of temperature and pressure; equal volume of different gases, contain the same number of moles (Avogadro's law).

Number of moles of oxygen = Number of moles of compound

$$\frac{\text{Mass of oxygen}}{\text{Molar mass of oxygen}} = \frac{\text{Mass of the compound}}{\text{Molar mass of the compound}}$$

$$\frac{\text{Molar mass of the compound}}{\text{Molar mass of oxygen}} = \frac{\text{Mass of the compound}}{\text{Mass of oxygen}}$$

But $\frac{\text{Mass of the compound}}{\text{Mass of oxygen}} = 5.25$

Thus $\frac{\text{Molar mass of the compound}}{\text{Molar mass of oxygen}} = 5.25;$

$$\text{Molar mass of the compound} = 5.25 \times \text{Molar mass of oxygen}$$

$$= 5.25 \times 32\text{g/mol} = 168\text{g/mol}$$

Let the molecular formula be $(C_3H_4O)_n$

Then $\big((3 \times 12) + (4 \times 1) + 16\big)n = 56n = M_r = 168$

$$n = \frac{168}{56} = 3$$

Then the molecular formula = $(C_3H_4O)_n = (C_3H_4O)_3 = C_9H_{12}O_3$

Hence the molecular formula is $C_9H_{12}O_3$

Example 12

A certain hydrocarbon is found to consist of 83.6% carbon and 16.4% hydrogen (by mass). If one molecule of this compound has a mass of 1.43×10^{-22}g, without finding empirical formula, calculate its molecular formula.

Solution

Given one molecule has mass of 1.43×10^{-22}g

Then using:

M_r = mass of one molecule (particle) $\times N_A$

$$= 1.43 \times 10^{-22}\text{g/molecule} \times 6.02 \times 10^{23}\text{molecule/mol}$$

$= 86\text{g/mol}$

Therefore, molar mass of the compound is 86g/mol

Constituent atom	**C**	**H**
Percentage composition	83.6	16.4
Mass of each in 86g (1mol) of the compound	72g	14g
Number of moles of each; $n = \frac{m}{M_r}$	$\frac{72\text{g}}{12\text{gmol}^{-1}} = 6\text{mol}$	$\frac{14\text{g}}{1\text{gmol}^{-1}} = 14\text{mol}$

Thus in one mole of the compound there are 6mol of carbon atoms and 14mol of hydrogen atoms and therefore in one molecule of the compound there are 6 atoms of carbon and 14 atoms of hydrogen.

Hence the molecular formula of the compound is C_6H_{14}.

Example 13

A 4.99g sample of compound contains 1.52g of nitrogen atoms and 3.47g of oxygen atoms. The molar mass of the compound is between 90g and 95g.

(i) Determine empirical and molecular formula
(ii) Calculate the actual molar mass of this compound.

Solution

Constituent atom	N	O
Mass of each in 4.99g of the compound	1.52g	3.47g
Number of mole of each; $n = \frac{m}{M_r}$	$\frac{1.52g}{14gmol^{-1}} = 0.1086mol$	$\frac{3.47g}{16gmol^{-1}} = 0.2169mol$
Divide by the smallest number to get simpler ratio	$\frac{0.1086mol}{0.1086mol} = 1$	$\frac{0.2169mol}{0.1086mol} = 2$

Thus the empirical formula is NO_2

Let Molecular formula be $(NO_2)_n$

Then $\left(14 + (2 \times 16)\right)n = 46n = M_r$

Using the lower limit of the $M_r(90\ g/mol)$

$$46n = 90; n = \frac{90}{46} = 1.957$$

Using upper limit of the $M_r(95g/mol)$

$$46n = 95; n = \frac{95}{46} = 2.0652$$

Thus the actual value of n must be the whole number between 1.957 and 2.0652 which is 2. Hence $n = 2$.

So the molecular formula = $(NO_2)_n$= $(NO_2)_2$= N_2O_4

Hence the molecular formula is N_2O_4

Since the molecular formula is N_2O_4;

Molar mass of $N_2O_4 = (2 \times 14) + (16 \times 4) = 92g/mol$

Hence:

(i) Empirical formula is NO_2 and molecular formula is N_2O_4
(ii) The actual molar mass is $92g/mol$

Example 14

A gaseous hydrocarbon, X, contains 85.7% of carbon by mass. 4.2g of the gas, X occupies a volume of $3.36dm^3$ at standard temperature and pressure.

(a) Determine the empirical formula
(b) What is the molecular formula of X?

Solution

'Hydrocarbon' means the compound consist of hydrogen and carbon only.

Thus with the given 85.7% of carbon;

$$\%\text{ of hydrogen} = 100 - \%\text{ of carbon} = 100 - 85.7 = 14.3\%$$

Constituent atom	**C**	**H**
Percentage composition	85.7	14.3
Mass of each in 100g of the compound	85.7g	14.3g
Number of mole of each; $n = \frac{m}{M_r}$	$\frac{85.7g}{12gmol^{-1}}$ $= 7.14mol$	$\frac{14.3g}{1gmol^{-1}}$ $= 14.3\ mol$
Dividing by smallest to get simpler ratio	$\frac{7.14\ mol}{7.14\ mol} = 1$	$\frac{14.3\ mol}{7.14\ mol} = 2$

Thus the empirical formula is CH_2.

Using $n = \frac{m}{M_r}$ and $n = \frac{V}{GMV}$

From which (by equating the two equations)

$$\frac{m}{M_r} = \frac{V}{GMV};$$

$$M_r = \frac{GMV \times m}{V} = \frac{22.4\ dm^3/mol \times 4.2g}{3.36\ dm^3} = 28g/mol$$

Therefore the relative molecular mass is $28g/mol$

Let molecular formula be $(CH_2)_n$

Then $\left(12 + (1 \times 2)\right)n = 28; 14n = 28; n = \frac{28}{14}; n = 2$

So the molecular formula $= (CH_2)_n = (CH_2)_2 = C_2H_4$

Hence the molecular formula is C_2H_4

PRACTICE EXERCISE 6

Question 1

The most common form of (Nylon – 6) is 63.68% carbon, 12.38% nitrogen, 9.8% hydrogen and 14.14% oxygen. Calculate the empirical formula for Nylon – 6.

Question 2

10g sample of compound contains 4g carbon, 0.667g hydrogen and 5.33 oxygen. Find the empirical and molecular formula of the compound (Molar mass of the compound is 180g/mol).

Question 3

A compound is 43.7% phosphorous and 56.3% oxygen by mass and has a vapour density of 142. Determine the molecular formula and hence empirical formula.

Question 4

An oxide of chromium is found to have the following percentage composition: 68.4% chromium and 31.6% oxygen. Determine the compound's empirical formula.

Question 5

When oleic acid is analyzed, it is found to consist of 76.6% carbon 12.1% hydrogen and 11.3% oxygen (by mass). It is also found that there are exactly 18 times as many atoms present in one molecule of oleic than there are in one molecule of water. Calculate the molecular formula of oleic acid.

Question 6

Determine the empirical formula of methane given that 6g of methane can be decomposed into 4.5g of carbon and 1.5g of hydrogen.

Question 7

Determine the empirical formula of the compound made when 8.65g of iron combines with 3.72g of oxygen (Molar mass of iron = 56)

Question 8

The composition of compound is 40% sulphur and 60% oxygen by weight. What is empirical formula?

Question 9

A sugar contains 39.95% carbon, 6.71% hydrogen and 53.34% oxygen by mass. If the molar mass of sugar was found experimentally to be 180g/mol; calculate its molecular formula.

Question 10

A 170g of an unidentified compound contains 29.84g sodium, 67.49g chromium and 72.67g oxygen. What is the compound's empirical formula?

Chapter seven

STOICHIOMETRIC CALCULATIONS

Stoichiometric calculation *is the calculation of the quantities of reactants and product in chemical reaction.*

Stoichiometric calculation used to study, relationship of amount of reactant used and product produced.

So if we have given amount of reactant; by stoichiometric calculation we can figure out what product is formed.

And if we have given amount of product formed we can figure out the amount of reactant used.

The reader should understand that in stoichiometric calculation comparison is done by using mole.

Below are steps used in solving stoichiometric calculations:

(i) Balancing given chemical equation
(ii) Convert the amount of a given substance to a mole
(iii) Using the mole ratio, calculate the mole of substance required
(iv) Convert the mole of the required substance into a form required, e.g. if you told to find mass, convert it to mass.

Example 1

Heating potassium chlorate, $KClO_3$, produces oxygen and potassium chloride. What mass of oxygen is obtained by quatitavely decomposing 3g of potassium chlorate?

Solution

Mass of potassium chlorate, m = 3g

Molar mass = $KClO_3 = 39 + 35.5 + (16 \times 3) = 122.5\text{g/mol}$

Number of moles, n = $\frac{m}{M_r} = \frac{3g}{122.5g/mol} = 0.02$ mol

So number of moles of potassium chlorate is0.02 mol

$2KClO_3 \rightarrow 3O_2 + 2KCl$

$2\text{mol of } KClO_3 \equiv 3 \text{ mol of } O_2$

$0.02\text{mol of } KClO_3 \equiv \frac{0.02 \text{ mol} \times 3 \text{ mol}}{2 \text{ mol}} = 0.03 \text{ mol of } O_2$

From $n = \frac{m}{M_r}$; $m = nM_r = 0.03\text{mol} \times 32\text{g/mol} = 0.96\text{g}$

Hence the mass of O_2 is 0.96g

Example 2

What mass of carbon dioxide and water is produced from combustion of 10g of lighter fluid butane(C_4H_{10}) and what mass of oxygen is consumed in the reaction?

Solution

Combustion of butane produces is carbon dioxide and water according to the following equation:

$$2C_4H_{10} + 13O_2 \rightarrow 8CO_2 + 10H_2O$$

Given:

Mass of $C_4H_{10} = 10\text{g}$

Molar mass of $C_4H_{10} = (12 \times 4) + (10 \times 1)$= 58g/mol

Then $n = \frac{m}{M_r} = \frac{10\text{g}}{58\text{g/mol}} = 0.17 \text{ mol}$

From the balanced chemical equation above:

$2\text{mol of } C_4H_{10} \equiv 8 \text{ mol of } CO_2$

$0.17\text{mol of } C_4H_{10} \equiv \frac{0.17 \times 8}{2}$ mol= 0.68 mol

From $n = \frac{m}{M_r}$; $m = nM_r$;where $M_r(CO_2) = 44\text{g/mol}$

Thus mass of $CO_2 = 0.68 \times 44\text{g} = 29.92\text{g}$

The mass of carbon dioxide = 29.92g

Also;

$$2\text{mol of } C_4H_{10} \equiv 10\text{mol of } H_2O$$

$$0.17\text{mol of } C_4H_{10} \equiv \frac{0.17 \times 10}{2} \text{ mol} = 0.85 \text{ mol of } H_2O$$

Then from$m = nM_r$; where $M_r(H_2O) = 18g/mol$

Mass of water $= 0.85mol \times 18g/mol = 15.3g$

Mass of H_2O is $15.3g$

Also;

$$2\text{mol of } C_4H_{10} \equiv 13\text{mol of } O_2$$

$$0.17\text{mol of } C_4H_{10} \equiv \frac{0.17 \times 13}{2} \text{mol} = 1.105 \text{ mol}$$

Then from $m = nM_r$;where $M_r(O_2) = 32g/mol$

Mass of oxygen $= 1.105mol \times 32g/mol$= 35.36g

Mass of oxygen consumed = 35.36g

Example 3

What mass ammonium sulphate is required to produce 3mol of ammonia from the following reaction:

$$(NH_4)_2SO_4 + Ca(OH)_2 \rightarrow CaSO_4 + NH_3 + H_2O$$

Solution

Here the equation is not balanced; so firstly balancing the equation:

$(NH_4)_2SO_4 + Ca(OH)_2 \rightarrow CaSO_4 + 2NH_3 + 2H_2O$(balanced)

$$2 \text{ mole of } NH_3 \equiv 1\text{mole of } (NH_4)_2SO_4$$

$$3 \text{ mole of } NH_3 \equiv \frac{3 \times 1}{2}\text{mole} = 1.5\text{mole of } (NH_4)_2SO_4$$

Mole of $(NH_4)_2SO_4 = 1.5 \text{ mol}$

From $n = \frac{m}{M_r}$; $m = nM_r$;where $M_r((NH_4)_2SO_4) = 132gmol^{-1}$

Mass of $(NH_4)_2SO_4 = 1.5\ mol\ \times 132g/mol = 198\ g$

Hence mass of ammonium sulphate is 198g

Example 4

What is the maximum number of grams of silver chloride that will precipitate from solution made by mixing solution of silver nitrate and 25mL of 0.05M $MgCl_2$?

Solution

Using $n = MV$;

Number of moles of $MgCl_2 = 25 \times 10^{-3} \times\ 0.05mol/L$

$$= 1.25 \times 10^{-3}mol$$

Maximum mass of AgCl is obtained when all given amount of $MgCl_2$ is converted to AgCl by its reaction with the silver nitrate according to the following equation:

$$2AgNO_3 +\ MgCl_2\ \rightarrow 2AgCl + Mg(NO_3)_2$$

From which mole ratio of $MgCl_2$ to AgCl is 1:2

And thus number of moles of AgCl produced

$$= 1.25 \times 10^{-3}mol\ \times 2 = 2.5 \times 10^{-3}mol$$

Then fromm $= n \times M_r$; where $M_r$of AgCl $= 143.5g/mol$

Substituting $m = 2.5 \times 10^{-3}mol\ \times 143.5g/mol = 0.3588g$

Hence mass of AgCl is 0.3588g

Example 5

Iron is produced when carbon monoxide reacts with iron (III) oxide according to the equation;

$$Fe_2O_{3(g)} + 3CO_{(g)}\ \rightarrow 2Fe_{(l)} + 3CO_{2(g)}$$

If 85tonnes of Iron was produced in such reaction; what mass of carbon dioxide would also be produced?

Solution

Given:

$Fe_2O_{3(g)} + 3CO_{(g)} \rightarrow 2Fe_{(l)} + 3CO_{2(g)}$

Number of mole of iron, n= $\frac{\text{mass}}{\text{molar mass}}$; where M_r of Fe is 56g/mol and mass= $85\text{tonnes} = 85 \times 10^6\text{g}$ $(1\text{ tonne} = 10^6\text{g})$

Substituting $n = \frac{85 \times 10^6\text{g}}{56\text{g/mol}} = 1517857\text{mol}$

But from the given equation, mole ratio of CO_2: Fe is 3:2;

That is $\frac{n_{CO_2}}{n_{Fe}} = \frac{3}{2}$; $\frac{n_{CO_2}}{1517857} = \frac{3}{2}$;

$$n_{CO_2} = \frac{3}{2} \times 1517857\text{mol} = 2276785.5\text{mol}$$

Then using $m = n \times M_r$; where $M_r \text{of} CO_2 = 44\text{g/mol}$

Mass of $CO_2 = 2276785.5\text{mol} \times 44\text{g/mol}$

$$= 100178562\text{g} = \frac{100178562}{10^6}\text{tonnes} = 100.178562\text{tonnes}$$

Mass of CO_2 produced is approximately 100tonnes.

Alternative solution

Below is the quicker method to solve this kind of problem.

But from the given reaction equation, mole ratio of CO_2: Fe is 3:2;

Then using $m = n \times M_r$;

$$3\text{mol of } CO_2 = 3 \times 44\text{g} = 132\text{g of } CO_2$$

And $2\text{mol of Fe} = 2 \times 56\text{g} = 112\text{g of Fe}$

Thus $112\text{g of Fe} \equiv 132\text{g of } CO_2$

And in tonnes; $112\text{tonnes of Fe} \equiv 132\text{tonnes of } CO_2$

Whence $85\text{tonnes of Fe} \equiv \frac{85\times132}{112} = 100\text{tonnes of } CO_2$

Hence mass of CO_2 is 100tonnes.

Example 6

6g of sample of potassium chlorate $(KClO_3)$ gave 1.9g of oxygen on strong heating. What is the percentage purity of the sample?

Solution

On heating, $KClO_3$ decompose according to the following equation:

$$2KClO_3 \rightarrow 2KCl + 3O_2$$

From which mole ratio of $KClO_3$ to O_2 is $2:3$

That is $\frac{n_{KClO_3}}{n_{O_2}} = \frac{2}{3}$ or $n_{KClO_3} = \frac{2}{3}n_{O_2}$

But from $n = \frac{m}{M_r}$;

$$n_{O_2} = \frac{\text{Mass of oxygen }(O_2)}{\text{Molar mass of oxygen }(O_2)}$$

$$= \frac{1.9\text{g}}{32\text{gmol}^{-1}} = 0.059375\text{mol}$$

Thus $n_{KClO_3} = \frac{2}{3}n_{O_2} = \frac{2}{3} \times 0.059375\text{mol} = 0.0396\text{mol}$

Also from $n = \frac{m}{M_r}$; $m = nM_r$

Thus mass of pure $KClO_3$ decomposed $= 0.0396 \times M_r(KClO_3)$

$$M_r(KClO_3) = \big(39 + 35.5 + (3 \times 16)\big)\text{gmol}^{-1} = 122.5\text{gmol}^{-1}$$

Therefore mass of pure $KClO_3 = 0.0396 \times 122.5\text{g} = 4.851\text{g}$

$$\% \text{ purity} = \frac{\text{mass of pure}(KClO_3)}{\text{mass of impure}} \times 100\%$$

$$= \frac{4.851}{6g} \times 100\% = 80.85\%$$

Hence the percentage purity of $KClO_3$ is 80.85%

Example 7

Find out the weight of $CaCO_3$ that must be decomposed to produce sufficient quantity of carbon dioxide to convert 10.6g of Na_2CO_3 completely into $NaHCO_3$.

Solution

$CaCO_3$ decomposes according to the following equation:

$$CaCO_3 \rightarrow CaO + CO_2$$

The produced CO_2 when dissolved in water, react with Na_2CO_3 to produce $NaHCO_3$ according to the following equation:

$$Na_2CO_3 + CO_2 + H_2O \rightarrow 2NaHCO_3$$

Using $n = \frac{m}{M_r}$; where $M_r(Na_2CO_3) = 106g/mol$

Number of moles of $Na_2CO_3 = \frac{10.6g}{106g/mol} = 0.1mol$

Mole ratio: $Na_2CO_3 : CO_2 = 1:1$

Hence number of moles of CO_2 is also 0.1mol

Also from the mole ratio: $CaCO_3 : CO_2 = 1:1$;

Number of moles of $CaCO_3$ was also 0.1mol

Then using $n = \frac{m}{M_r}$; where $M_r(CaCO_3) = 100g/mol$

Mass of $CaCO_3 = 0.1 \text{ mol} \times 100g/mol = 10g$

Hence the weight of $CaCO_3$ is 10g

Example 8

12.46g of a mixture of MgO and $MgCO_3$ on strong heating lost 4.4g in weight. What is the composition of mixture?

Solution

When the mixture is heated only, $MgCO_3$ will decompose according to the following equation:

$$MgCO_3 \rightarrow MgO + CO_2$$

Weight loss = Weight of CO_2 formed

Therefore mass of CO_2 is 4.4g

Using $n = \frac{m}{M_r}$; where $M_r(CO_2) = 44g/mol$

Number of molesCO_2 formed $= \frac{4.4g}{44\ gmol^{-1}} = 0.1\ mol$

But from the equation of decomposition of $MgCO_3$, mole ratio of $MgCO_3$ to CO_2 is $1:1$

Thus $\frac{n_{MgCO_3}}{n_{CO_2}} = \frac{1}{1}$ or $n_{MgCO_3} = n_{CO_2}$

So n_{MgCO_3} was also 0.1mol

Also using $m = nM_r$;

Mass of $MgCO_3$ in the mixture = $0.1 \times M_r(MgCO_3)$

Where$M_r(MgCO_3)$= 84g/mol

Thus mass of $MgCO_3$ in the mixture = $0.1mol \times 84g/mol = 8.4g$

The $\%MgCO_3 = \frac{\text{Mass of } MgCO_3}{\text{Mass of mixture}} \times 100\%$

$$= \frac{8.4}{12.46} \times 100\% = 67.4\%$$

And $\%\ MgO = (100 - 67.4)\%$= 32.6%

Hence the mixture is 67.4% $MgCO_3$ and 32.6% MgO

Example9

A mixture of cuprous oxide (Cu_2O) and cupricoxide (CuO) was found to contain 88% copper. Calculate the amount of each oxide in 2g of sample of the mixture.

Solution

Let mass of Cu_2O in the mixture be x in grams

And mass of CuO in the mixture be y in grams

Then $x + y = 2$ ……………………………..(i)

Then molar mass of $Cu_2O = \big((2 \times 64) + 16\big)\text{g/mol} = 144\text{g/mol}$

And molar mass of $CuO = (64 + 16)\text{g/mol} = 80\text{g/mol}$

In one mole of Cu_2O : $144\text{g of } Cu_2O \equiv 128\text{g of Cu}$

In xg of Cu_2O : $\text{xg of } Cu_2O \equiv ?$

$$144\text{g}(Cu_2O) \equiv 128\text{g (Cu)}$$
$$\text{xg}(Cu_2O) \equiv \text{Mass of Cu in xg of } Cu_2O$$

Mass of Cu in xg of $Cu_2O \times 144 = 128x$

Hence mass of Cu in xg of $Cu_2O = \frac{128}{144}x$

Also in one mole of CuO : $80\text{g of CuO} \equiv 64\text{g of Cu}$

In yg of CuO : .yg of CuO $\equiv$?

Again by cross multiplication;

Mass of Cu in yg of $CuO = \frac{64}{80}y$

But total mass of Cu in 2g of the mixture

= Mass of Cu in Cu_2O + Mass of Cu in $CuO = \frac{128}{144}x + \frac{64}{80}y$

And % Cu $= \frac{\text{Total mass of Cu}}{\text{Mass of the mixture}} \times 100\%$

Then $88 = \left(\frac{\frac{128}{144}x + \frac{64}{80}y}{2}\right) \times 100$

From which;

$$\frac{128x}{144} + \frac{64y}{80} = 1.76 \quad \ldots\ldots\ldots\ldots\ldots\ldots\ldots\ldots\ldots..(ii)$$

Solving equation (i) and (ii) simultaneously gives;

$$x = 1.8 \qquad y = 0.2$$

Hence:

Mass of Cu_2O in the mixture is 1.8g

Mass of CuO in the mixture is 0.2g

Example 10

A 1.85g sample of mixture of $CuCl_2$ and $CuBr_2$ was dissolved in water and mixed thoroughly with 1.8g portion of AgCl. After reaction, the solid which now contains AgCl and AgBr was filtered, dried and weighted to be 2.052g. What is percentage by weight of $CuBr_2$ in the mixture?

Solution

When the mixture of $CuBr_2$ and $CuCl_2$ is mixed with AgCl, only $CuBr_2$ will react with AgCl to give AgBr according to the following equation:

$$CuBr_2 + 2AgCl \rightarrow CuCl_2 + 2AgBr$$

From which mole ratio of AgCl to AgBr is $1:1$ $(2:2)$

That is $\frac{n_{AgCl}}{n_{AgBr}} = \frac{1}{1}$ or $n_{AgCl} = n_{AgBr}$

Presence of AgCl after the above reaction, implies that AgCl was in excess and some of it remained unreacted after the reaction.

Now, let mass of reacted AgCl be x in grams

And mass of unreacted AgCl be y in grams

The using $n = \frac{m}{M_r}$;

Number of mole of reacted AgCl $= \frac{x}{143.5}$

And as $n_{AgCl} = n_{AgBr}$: number of moles of produced AgBr, n_{AgBr} was also $\frac{x}{143.5}$

Also using; $m = nM_r$;

Mass of AgBr produced in the reaction $= \frac{x}{143.5} \times 188 = \frac{1880x}{1435}$

But after the reaction;

Total mass of the solid

= mass of produced AgBr + mass of unreacted AgCl $= 2.052g$

That is $\frac{1880x}{1435} + y = 2.052$ ....(i)

But also total mass of AgCl was 1.8g

That is $x + y = 1.8$ ....(ii)

Taking (ii) – (i) gives;

$$\frac{1880x}{1435} - x = 2.052 - 1.8$$

$$\frac{89x}{287} = 0.252; x = \frac{0.252 \times 287}{89} = 0.8126$$

Thus mass of reacted AgCl $= 0.8126g$

And number of moles of reacted AgCl $= \frac{0.8126g}{143.5g/mol}\left(n = \frac{m}{M_r}\right)$

$= 5.6627 \times 10^{-3}$mol

Also from the reaction equation;

$$CuBr_2 + 2AgCl \rightarrow CuCl_2 + 2AgBr$$

Mole ratio of $CuBr_2$ to AgCl is $1:2$

That is $\frac{n_{CuBr_2}}{n_{AgCl}} = \frac{1}{2}$

From which $n_{CuBr_2} = \frac{1}{2} \times n_{AgCl} = {}^1/_2 \times 5.6627 \times 10^{-3}mol$

$= 2.83135 \times 10^{-3}mol$

Then using $m = nM_r$;

Mass of $CuBr_2 = 2.83135 \times 10^{-3} \times 224g = 0.6342g$

And % $CuBr_2 = \frac{\text{mass of } CuBr_2}{\text{mas of the mixture}} \times 100\%$

$= \frac{0.6342}{1.85} \times 100\% = 34.28\%$

Hence the percentage of $CuBr_2$ is 34.28%

LIMITED AND EXCESS REACTANTS

The chemical reactions we have worked so far, have assumed that reactants are present in the exact mole ratio as appear in the balanced equation. In other words, we did not have to worry about one or more reactants remaining out after the reaction was complete.

However, when reactions are carried out in the laboratory, we do not always add reactants in the exact mole ratio as indicated by the balanced equation. For various reasons, one or more reactant is usually present in excess.

So in most cases, only one reactant is completely consumed at the end of reaction and this reactant limits (determines) the amount of products that can be formed and hence it is known as **limited reactant** (or **regent**) while another reactant that present in excess is known as **excess reactant** (or **reagent**). Once the limiting reagent is entirely consumed the reaction stops, the other excess reactants will still be present in some amount.

By definition:

Limiting reactant *is the reactant that is completely used up during the chemical reaction.*

Excess reactant *is the reactant that is not completely used up during the chemical reaction.*

To have deeper understanding of the concept, consider the following hypothetical reaction between A and B as per the following equation:

$$2\mathrm{A} + \mathrm{B} \longrightarrow \mathrm{C}$$

To make things more understandable and interesting, imagine A represents a spoon of sugar, B represents a cup of hot water and C represents a cup of tea such that (of course with other ingredients):

$$\text{Two spoons of sugar} + \text{One cup of hot water} \longrightarrow \text{One cup of tea}$$

Then consider the following three different cases:

Case 1: $n_A = n_B$ e.g. 2 moles of A (two spoons of sugar) are mixed with 2 moles of B (two cups of hot water).

In this case B is excess reactant while A is limited reactant because from the stoichiometry of the equation, 2 moles of A needs only 1 mole of B; that is to prepare one cup of tea (C), only one cup of hot water (B) needs two spoons of sugar (A) while the given two cups of hot water (2 moles of B) needs four spoons of sugar (4 moles of A).

So no matter how much hot water is present, the amount of product, C (tea) will be limited (determined) by the amount of sugar (A) and in our example, a maximum of one cup of tea (1 mole of C) will be produced and 1 mole of B will remain unreacted after the reaction.

To summarize; A is limited reactant because $\mathrm{n_A} < 2\mathrm{n_B}$.

Case 2: $\mathbf{n_A > n_B}$

Here there are more three possibilities:

First possibility: $\mathbf{n_A > n_B}$ and $\mathbf{n_A = 2n_B}$ e.g. 4 moles of A (four spoons of sugar) are mixed with 2 moles of B (two cups of hot water).

Here the given amount satisfies mole ratio, no reagent presents in excess and all reactants (reagents) will completely be consumed in the reaction to give 2 moles of product, C (two cups of tea).

Second possibility: $\mathbf{n_A > n_B}$ and $\mathbf{n_A < 2n_B}$ e.g. 3 moles of A (three spoons of sugar) are mixed with 2 moles of B (two cups of hot water).

Here A will be limited reactant although its number of moles is greater than that of B because 2 moles of B (two cups of hot water) needs 4 moles of A (four spoons of sugar) while only 3 moles of A (three spoons of sugar) are present.

Hence the amount of product will be determined by A whereby 3 moles of A will react with 1.5 moles of B to give 1.5 moles of C and 0.5 mole of B will remain unreacted.

Third possibility: $\mathbf{n_A > n_B}$ and $\mathbf{n_A > 2n_B}$ e.g. 4 moles of A (four spoons of sugar) are mixed with 1 mole of B (one cup of hot water).

Here it is clear that, B is limited reactant because 1 mole of B needs only 2 moles of A and hence 2 moles of A will remain unreacted at the end of the reaction.

Case 3: $\mathbf{n_A < n_B}$ e.g. 2 moles of A (two spoons of sugar) are mixed with 3 moles of B (three cups of hot water).

Here it is clear that A is limited reactant because the equation suggests that n_A must be twice of n_B (number of spoons of tea must double the number of cups of hot water).

So because $n_A < n_B$which obvious implies that $n_A < 2n_B$, A is limited reactant.

Example 11

A piece of iron with mass 5.59g is ignited in a vessel containing 1.6g oxygen to form Fe_3O_4. Deduce which reactant is in excess and calculate the amount in moles by which it is in excess.

Solution

Given:

Mas of Iron, $m = 5.59g$

Mass of Oxygen = 1.6g

$3Fe + 2O_2 \quad \rightarrow \quad Fe_3O_4$

Mole ratio;Fe : $O_2 = 3:2$

$$n_{Fe} = \frac{m_{Fe}}{M_r(Fe)} = \frac{5.59g}{56g/mol} = 0.0998mol$$

$$n_{O_2} = \frac{m_{O_2}}{M_r(O_2)} = \frac{1.6g}{32g/mol} = 0.05mol$$

$$0.0998\text{mol of Fe} \times \frac{2\text{mol of }O_2}{3\text{ mol of Fe}} = 0.0665\text{mol of }O_2$$

$$0.05\text{mol of }O_2 \times \frac{3\text{ mol of Fe}}{2\text{ mol of }O_2} = 0.075\text{mol of Fe}$$

Since 0.075mol of Fe but we have0.0998mol of Fe, iron is excess reactant

Number of mole of excess = $(0.0998 - 0.075)$mol= 0.025mol

Therefore number of moles of Fe in excess is 0.025mol

Example 12

Carbon with mass of 1.2g is completely burned in a closed container with 8g of oxygen.

(a) What mass of carbon dioxide would be formed?
(b) Which reagent is in excess and what mass of it remain after reaction?

Solution

Balanced equation for the reaction is:

$$C + O_2 \rightarrow CO_2$$

Number of mole of C = $\frac{\text{Mass of C}}{\text{Molar mass of C}}$

$$= \frac{1.2\text{g}}{12\text{g/mol}} = 0.1\text{mol}$$

Number of mole of $O_2 = \frac{\text{Mass of } O_2}{\text{Molar mass of } O_2}$

$= \frac{8\text{g}}{32\text{g/mol}} = 0.25\text{mol}$

Mole ratio;$C : O_2 = 1:1$

So 0.1 mol of C need 0.1mol of O_2

Thus carbon is limiting reactant

$$1 \text{ mol of C} \equiv 1 \text{ mol of } CO_2$$

$$0.1\text{mol of C} \equiv \frac{0.1 \times 1}{1} = 0.1\text{mol of } CO_2$$

From $n = \frac{m}{M_r}$; $m = n \times M_r$; where$M_r(CO_2) = 44\text{g/mol}$

Thus mass of $CO_2 = 0.1\text{mol} \times \frac{44\text{g}}{\text{mol}} = 4.4\text{g}$

Hence mass of carbon dioxide formed is 4.4g

(a) Oxygen gas is excess reactant.

Total number of moles of O_2 = 0.25mol

Number of moles of O_2reacted = 0.1mol

Number of moles of unreacted O_2 = $(0.25 - 0.1)$mol= 0.15 mol

Using $m = n \times M_r$; where M_r of $O_2 = 32g/mol$

$$m = 0.15 \text{ mol} \times 32g/mol = 4.8g$$

Hence mass of oxygen remained is 4.8g

Example 13

What weight of AgCl will be formed when solution of containing 4.77g of NaCl is added to a solution of 5.77g of $AgNO_3$

Solution

$NaCl + AgNO_3 \rightarrow AgCl + NaNO_3$

Molar mass of $NaCl = (23 + 35.5)g/mol = 58.5g/mol$

Number of mole of $NaCl = \frac{m}{M_r} = \frac{4.77g}{58.5g/mol} = 0.0815mol$

Molar mass of $AgNO_3 = \big(108 + 14 + (16 \times 3)\big)g/mol$

$$= 170g/mol$$

Number of mole of $AgNO_3 = \frac{m}{M_r} = \frac{5.77g}{170g/mol} = 0.0339mol$

From the reaction equation; $NaCl + AgNO_3 \rightarrow AgCl + NaNO_3$

Mole ratio of $NaCl: AgNO_3 = 1:1$

Thus $0.0339mol$ of $AgNO_3$ needed only $0.0339mol$ of NaCl

and hence $AgNO_3$ is limited reactant while NaCl is excess reactant.

Also from the same reaction equation above; mole ratio of $AgNO_3$ (limited reactant) to AgCl is 1: 1. Thus number of moles of AgCl formed was also 0.0339mol.

Then using $m = n \times M_r$; where$M_r$of $AgCl = 143.5g/mol$

$m_{AgCl} = 0.0339 \times 143.5 = 4.86g$

Hence mass of AgCl formed is 4.86g.

Example 14

100mL of 1M KOH is mixed with 32.5mL of 2M H_3PO_4 and allowed to react according to the equation:

$3KOH_{(aq)} + H_3PO_{4(aq)} \rightarrow K_3PO_{4(aq)} + 3H_2O$

Calculate the mass of K_3PO_4 produced.

Solution

Number of moles of $KOH = M \times V = 1 \times 100 \times 10^{-3} = 0.1mol$

Number of mole of H_3PO_4 = M $\times$ V= 2 $\times$ 32.5 $\times 10^{-3}$= 0.065mol

But from the given reaction equation, mole ratio of KOH to H_3PO_4 is 3:1; that is 0.065mol of H_3PO_4 needs $3 \times 0.065 = 0.195$mol of KOH which is greater than the given amount. Thus KOH is limited reactant and is the one which will determine amount of K_3PO_4 produced.

Then from;

$3KOH_{(aq)} + H_3PO_{4(aq)} \rightarrow K_3PO_{4(aq)} + 3H_2O$

Mole ratio of KOH to K_3PO_4 is 3: 1;

Thus number of moles of $K_3PO_4 = \frac{0.1 \times 1}{3} = 0.033mol$

Then using $m = n \times M_r$; where $M_r$of $K_3PO_4 = 212g/mol$

Mass of $K_3PO_4 = 0.033mol \times 212g/mol = 6.996g$

Hence mass of K_3PO_4 produced is 6.996g

PRACTICE EXERCISE 7

Question 1

Sodium hydroxide reacts with carbon dioxide to produce sodium carbonate and water. Calculate the mass of sodium hydroxide required to prepare 53g of sodium carbonate.

Question 2

Hydrogen and chlorine gas combine to form hydrogen chloride gas, HCl. Calculate the minimum mass of hydrogen gas and chlorine gas needed to produce 10g of hydrogen chloride.

Question 3

A piece of pure zinc weighing 15.8g dissolved completely in hydrochloric acid: Calculate:

(a) Mass of hydrogen gas formed.
(b) The mass of zinc chloride that could be formed by evaporating the water from the resulting solution.

Question 4

5g of sample of calcium carbonate $(CaCO_3)$ contaminated with some volatile impurity left a residue of 2.2g on strong heating. What is the percentage of pure $CaCO_3$ in the sample?

Question 5

1.42g of a mixture of $CaCO_3$ and $MgCO_3$ were heated till no further loss in weight takes place. The residue left was weighted and found to be 0.76g. Find the percentage composition of the mixture.

Question 6

How much quantity of zinc will have to be reacted with excess dilute HCl solution to produce sufficient hydrogen gas for completely reacting with oxygen obtained by decomposing 5.104g of potassium chlorate?

Question 7

Carbon dioxide can be prepared in the laboratory by reacting magnesium carbonate with dilute hydrochloric acid according to the following equation.

$$MgCO_{3(s)} + 2HCl_{(aq)} \rightarrow MgCl_{2(aq)} + H_2O_{(l)} + CO_{2(g)}$$

If 10g of CO_2 is produced, calculate:

(a) The mass of magnesium carbonate reacted.
(b) The mass of hydrochloric acid reacted.

Question 8

Consider the reaction exposed by this equation

$$Ag_{(s)} + 2HNO_{3(aq)} \rightarrow AgNO_{3(aq)} + H_2O + NO_{2(g)}$$

(a) What mass of Ag reacts with $28.5cm^3$ of $0.564M$ HNO_3?
(b) What mass of NO_2 is formed when $14.3cm^3$ of $2M$ HNO_3 reacted completely with Ag?
(c) What volume of $2M$ HNO_3 is needed to react completely with 120g of Ag?

Question 9

18g calcium oxide is mixed with 12g of carbon and allowed to react according to the equation:

$$CaO_{(s)} + 3C_{(s)} \rightarrow CaC_{2(s)} + CO_{(g)}$$

What mass of calcium carbide (CaC_2) is produced?

Question 10

Determine the amount of magnesium required to liberate $900cm^3$ of hydrogen from a solution of HCl at 27°C and 740mmHg.

Chapter eight

CALCULATIONS INVOLVING ACID – BASE TITRATION

INTRODUCTION

An **acid-base titration** *is the determination of the concentration of an acid or base by exactly neutralizing the acid/base with an acid or base of known concentration.*

- It makes use of **neutralization reaction** that occurs between acids and bases and the knowledge of how acids and bases will react if their formulas are known.

The method of quantitative analysis which involves the experimental determination of volumes of solutions which react together completely is known as **titrimetric analysis**.

Some important definitions related to acid-base titration

Standard solution *is the solution of known concentration in the titration.* It is also known as **titrant** or **titrator**. An aqueous sodium carbonate is a good example of titrant which is commonly used in the experiment.

Analyte or **titrand** *the solution of unknown concentration in the titration.*

Alkalimetry, (sometimes spelled **alkimetry**), *is the specialized analytic use of acid – base titration to determine the concentration of a basic substance.* Thus in alkalimetry, the analyte is the basic substance.

Acidimetry, (sometimes spelled **acidometry**), *is the specialized analytic use of acid – base titration to determine the concentration of an acidic substance.* Thus in acidimetry, the analyte is the acidic substance.

Uses of acid – base titration

Acid – base titration is useful in:

- Determination of unknown concentration of acid or base
- Determination of percentage purity of chemicals

Determination of unknown concentration of acid or base

If acid and base react according to the following equation:

$$\mathrm{n_aACID + n_bBASE \longrightarrow SALT + WATER}$$

Where $\mathrm{n_a}$ and $\mathrm{n_b}$ are stoichiometric coefficient for acid and base respectively in the balanced equation.

Thence mole ratio of acid to base will be $\mathrm{n_a : n_b}$

That is $\dfrac{\textbf{Number of moles of acid (used in the titration)}}{\textbf{Number of moles of base (used in the titration)}} = \dfrac{\mathbf{n_a}}{\mathbf{n_b}}$

If $\mathrm{M_a}$ and $\mathrm{M_b}$ represent molarity of acid and base respectively

And $\mathrm{V_a}$ and $\mathrm{V_b}$ represent volume of acid and base respectively

Then using $\mathrm{n = MV}$ $\left(\mathrm{From\ M = \frac{n}{V}}\right)$

$\text{Number of moles of acid (used in the titration)} = \mathrm{M_aV_a}$

$\text{Number of moles of base (used in the titration)} = \mathrm{M_bV_b}$

Hence $\dfrac{\mathbf{M_aV_a}}{\mathbf{M_bV_b}} = \dfrac{\mathbf{n_a}}{\mathbf{n_b}}$

The above final formula is very useful in calculating unknown concentration of either acid or base.

Determination of percentage purity of chemicals

This can be done with the aid of the following formula:

$$\textbf{Percentage purity} = \frac{\textbf{Mass of pure (in given volume of the solution)}}{\textbf{Mass of impure(in the same volume of solution)}} \times \mathbf{100}$$

If the volume of the solution is unit (e.g. $1dm^3$), then the mass in the given volume becomes equivalent to mass concentration and therefore the above formula becomes;

$$\textbf{Percentage purity} = \frac{\textbf{Mass concentration of pure}}{\textbf{Mass concentration of impure}} \times \mathbf{100}$$

Example 1

$15cm^3$ of $0.5MNaOH$ is used to neutralize $25cm^3$ of an acetic acid, CH_3COOH. What is the molar concentration of the acetic acid?

Solution

Given that:

Volume of base, $V_b = 15cm^3$

Molarity of base, $M_b = 0.5M$

Volume of acid, $V_a = 25cm^3$

$$CH_3COOH + NaOH \rightarrow CH_3COONa + H_2O$$

$$n_a = n_{b`} = 1$$

Thus from $\frac{M_aV_a}{M_bV_b} = \frac{n_a}{n_{b`}}$; $M_aV_a = M_bV_b$

$$M_a = \frac{M_bV_b}{V_a} = \frac{0.5M \times 15cm^3}{25cm^3} = 0.3M$$

Hence the molar concentration of the acid is 0.3M

Example 2

How many cm^3 of 0.2M KOH will exactly neutralize $15cm^3$ of 0.4M H_2SO_4?

Solution

Given that:

Molarity of base, $M_b = 0.2M$

Volume of acid, $V_a = 15cm^3$

Molarity of acid, $M_a = 0.4M$

Equation for the reaction;

$$H_2SO_4 + 2KOH \longrightarrow K_2SO_4 + 2H_2O$$

$n_a = 1, \; n_b = 2$

From$\frac{M_aV_a}{M_bV_b} = \frac{n_a}{n_b}$; $V_b = \frac{n_bM_aV_a}{n_aV_a} = \frac{2 \times 0.4M \times 15cm^3}{1 \times 0.2M} = 60cm^3$

Volume of base, is$60cm^3$

Example 3

How many millilitres of 2M sodium hydroxide are necessary to neutralize 25.00mL of 1.5M phosphoric acid?

Solution

Equation for the reaction;

$3NaOH + H_3PO_4 \rightarrow Na_3PO_4 + 3H_2O$

From which $n_a = 1, n_b = 3$

Then using $\frac{M_aV_a}{M_bV_b} = \frac{n_a}{n_b}$; Where $M_a = 1.5M, M_b = 2M, V_a = 25mL$

Substituting $\frac{1.5M\times25mL}{2M\times V_b} = \frac{1}{3}$; $V_b = \frac{3\times1.5\times25mL}{2\times1} = 56.25mL$

Hence volume of NaOH needed is 56.25mL

Example 4

What volume of 6MHCl is necessary to neutralize and dissolve 31.564g of solid aluminium hydroxide.

Solution

Using $n = \frac{m}{M_r}$; where M_r of $Al(OH)_3 = 78g/mol$

Number of moles of $Al(OH)_3 = \frac{\text{mass of } Al(OH)_3}{\text{molar mass of } Al(OH)_3}$

$$= \frac{31.564g}{78g/mol} = 0.4mol$$

From the reaction equation;

$3HCl + Al(OH)_3 \rightarrow AlCl_3 + 3H_2O$; $n_a = 3, n_b = 1$

Then using $\frac{\text{number of moles of acid}}{\text{number of moles of base}} = \frac{n_a}{n_b}$;

$\frac{\text{Number of moles of acid}}{0.4} = \frac{3}{1}$;

Number of moles of acid $= 3 \times 0.4mol = 1.2mol$

Then from $M = \frac{n}{V}$; $V = \frac{n}{M}$

Substituting $V = \frac{1.2mol}{6moldm^{-3}} = 0.2dm^3$

Hence the volume of HCl needed is $0.2dm^3$

Example 5

When 0.125g of diprotic organic acid is dissolved in water, it requires $17.5cm^3$ of 0.0893M KOH for neutralization. What is the molar mass of acid?

Solution

Mass of diprotic acid = 0.125g

Number of moles of KOH (base) = Molarity $\times$ Volume

$= 0.0893mol/dm^3 \times 17.5 \times 10^{-3}dm^3 = 1.563 \times 10^{-3}$ mol

Since the acid is diprotic (acid with from of H_2A), $n_b = 2$

Thus $n_a : n_b = 1:2; \frac{n_a}{n_b} = \frac{1}{2}$

Then from $\frac{\text{Number of moles of acid (used in the titration)}}{\text{Number of moles of base (used in the titration)}} = \frac{n_a}{n_b}$

Substituting $\frac{\text{Number of moles of acid}}{1.563 \times 10^{-3}\ \text{mol}} = \frac{1}{2}$

$$\text{From which number of moles of acid} = \frac{1.563 \times 10^{-3}\ \text{mol}}{2}$$

$$= 7.815 \times 10^{-4}\text{mol}$$

From $n = \frac{m}{M_r}$; $M_r = \frac{m}{n} = \frac{0.125\text{g}}{1.563 \times 10^{-3}\text{mol}} = 160\text{g/mol}$

Hence the molar mass is 160g/mol

Example6

18cm^3 of 0.05M H_2SO_4 is neutralized by 20cm^3 of a NaOH solution. On the other hand, 10cm^3 of oxalic acid is required to neutralize the same volume of NaOH solution. Determine the mass of oxalic acid crystals ($H_2C_2O_4 . 2H_2O$) used.

Solution

Equation for reaction between H_2SO_4 and NaOH;

$$H_2SO_4 + 2NaOH \rightarrow Na_2SO_4 + H_2O;\ n_a = 1, n_b = 2$$

Using $\frac{M_aV_a}{M_bV_b} = \frac{n_a}{n_b}$; $M_b = \frac{M_aV_an_b}{V_bn_a}$

$$= \frac{0.05 \times 18 \times 2}{20 \times 1}\text{M} = 0.09\text{M}$$

Equation for reaction between $H_2C_2O_4 . 2H_2O$ and NaOH

$H_2C_2O_4 . 2H_2O + 2NaOH \rightarrow Na_2C_2O_4 + 4H_2O$; $n_b = 2, n_a = 1$

Again using $\frac{M_aV_a}{M_bV_b} = \frac{n_a}{n_b}$; $M_a = \frac{M_bV_bn_a}{V_an_b}$

Where:

$M_b = 0.09M;\ V_b = 20cm^3$(Volume remained the same)

$V_a = 10cm^3$

Substituting $M_a = \frac{0.09 \times 20 \times 1}{10 \times 2}M = 0.09M$

Using $n = MV$;

Number of moles of oxalic acid crystals in $10cm^3$ of its solution $= \frac{10}{1000} \times 0.09mol = 9 \times 10^{-4}mol$

Using $m = n\ M_r$;

Mass of oxalic acid crystals = $9 \times 10^{-4} \times 126g = 0.1134g$

Example 7

Calculate mass of $Ca(OH)_2$ present in a sample if its treated to its equivalent point with $44.02cm^3$ of $0.0885MHNO_3$.

Solution

$Ca(OH)_2 + 2HNO_3 \rightarrow Ca(NO_3)_2 + 2H_2O$

Number of moles of $HNO_3 = 44.02 \times 10^{-3} \times 0.0885mol$

$= 3.896 \times 10^{-3}mol$

$2mol\ of\ HNO_3 \rightarrow 1\ mole\ of\ Ca(OH)_2$

$$3.896 \times 10^{-3}\ mol\ of\ HNO_3 \rightarrow \frac{3.896 \times 10^{-3}}{2}\ of\ Ca(OH)_2$$

$= 1.948 \times 10^{-3}\ mol$

Molar mass of $Ca(OH)_2 = 40 + 2 \times (16 + 1) = 74g/mol$

$n = \frac{m}{M_r}$; $m = n \times M_r = 1.948 \times 10^{-3} \times 74g = 0.144g$

The mass of $Ca(OH)_2$ is $0.144g$

Example 8

1.400g of pure anhydrous sodium carbonate was made up into $250cm^3$ of aqueous solution. $25.0cm^3$ of this solution required $24.50cm^3$ of a certain sample of hydrochloric acid.

(a) Calculate the following:

(i) Molarity of the acid

(ii) Concentration of acid ingdm^{-3}.

(b) If the remaining acid occupied $920cm^3$, how would it made exactly 0.1M?

(c) Is the titration acidimetry or alkalimetry?

Solution

Using mass concentration $= \frac{\text{Mass of solute}}{\text{Volume of solution}}$

$$\text{Mass concentration of } Na_2CO_3 = \frac{1.4g}{0.25dm^3} = 5.6gdm^{-3}$$

Then from molarity $= \frac{\text{mass concentration in gdm}^{-3}}{\text{Molar mass}}$

Where molar mass of $Na_2CO_3 = 106gmol^{-1}$

$$\text{Molarity of } Na_2CO_3, M_b = \frac{\text{mass concentration in gdm}^{-3}}{\text{Molar mass}}$$

$$= \frac{5.6gdm^{-3}}{106gmol^{-1}} = 0.05283M$$

Na_2CO_3 reacts with HCl according to the following equation:

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

From which $n_a = 2; n_b = 1$

Then using $\frac{M_aV_a}{M_bV_b} = \frac{n_a}{n_b}$;

$$M_a = \frac{M_b V_b n_a}{V_a n_{b`}} = \frac{0.05283M \times 25cm^3 \times 2}{24.5cm^3 \times 1} = 0.1078M$$

Molarity of the acid is 0.1078M

$\text{Mass concentration} = \text{Molarity} \times \text{Molar mass}$

= $0.1078 \times 36.5gdm^{-3} = 3.9347gdm^{-3}$

Thus mass concentration of the acid is $3.9347gdm^{-3}$

By dilution principle: $M_cV_c = M_dV_d$

Where: M_c is the molarity of concentrated solution = $0.1078M$

M_d is the molarity of diluted solution = 0.1M

V_c is the volume of concentrated solution = 920cm^3

V_d is the volume of diluted solution

Then $V_d = \frac{M_cV_c}{M_d} = \frac{0.1078 \times 920}{0.1}cm^3 = 991.8cm^3$

Volume of distilled water added = $V_d - V_c$

$= (991.8 - 920)cm^3 = 71.8cm^3$

Hence 71.8cm^3 of distilled water must be added to 920cm^3 of the acid solution in order to convert it to exactly 0.1M.

The titration is **acidimetry**

Example 9

3.000g of impure sodium carbonate were made up to 250cm^3 of aqueous solution. 25cm^3 of this solution required 21.00cm^3 of 0.1050M hydrochloric acid for its complete neutralization. Calculate the percentage purity of sodium carbonate.

Solution

Usingmass concentration $= \frac{\text{Mass of solute}}{\text{Volume of solution}}$

Mass concentration of impure $Na_2CO_3 = \frac{3g}{0.25dm^3} = 12gdm^{-3}$

Na_2CO_3 reacts with HCl according to the following equation:

$$Na_2CO_3 \quad + \quad 2HCl \quad \rightarrow \quad 2NaCl \quad + \quad CO_2 \quad + \quad H_2O$$

From which$n_a = 2; n_b = 1$

Then using $\frac{M_aV_a}{M_bV_b} = \frac{n_a}{n_b};$

Where $V_a = 21cm^3$, $M_a = 0.1050M$ and $V_b = 25cm^3$

Thence $M_b = \frac{M_aV_an_b}{V_bn_a} = \frac{0.1050M\times21cm^3\times1}{25cm^3\times2} = 0.0441M$

Then frommolarity $= \frac{\text{mass concentration in gdm}^{-3}}{\text{Molar mass}}$

Where molar mass of $Na_2CO_3 = 106gmol^{-1}$

Mass concentration (in gdm^{-3})of pure Na_2CO_3

$= \text{molarity} \times \text{molar mass} = 0.0441 \times 106 = 4.6746gdm^{-3}$

Then %Purity $= \frac{\text{mass concentration of pure}}{\text{mass of concentration of impure}} \times 100\%$

$= \frac{4.6746}{12} \times 100\% = 38.955\%.$

Hence percentage purity of sodium carbonate is38.955%

PRACTICE EXERCISE 8

Question 1

If $75cm^3$ of $0.823 M HClO_4$ requires $95.5cm^3$ of $Ba(OH)_2$ for complete neutralization, what is the molar concentration of the $Ba(OH)_2$ solution?

Question 2

If $26.4cm^3$ of LiOH solution are required to neutralize $21.7cm^3$ of $0.5 M HBr$. What is the molar concentration of the basic solution?

Question 3

Calculate the mass of $H_2C_2O_4$ present in a sample of it titrated to its equivalent point with $189.09cm^3$ of $0.2235M$ NaOH.

Question 4

What volume of $0.25M$ H_2SO_4 is needed to react completely with $13.5g$ of NaOH?

Question 5

22.4mL solution of HCl reacts with 20mL of $0.2M$ K_2CO_3. Determine mass and molar concentration of HCl.

Question 6

What mass of magnesium hydroxide powder (active ingredient in many anti-acid tablets) would be required to relieve (neutralize) an acid stomach if the acid in the stomach is 0.1M HCl and volume of the acid in the stomach is 800mL?

Question 7

8.58g of impure sodium carbonate were made up to $250cm^3$ of aqueous solution. $25cm^3$ of this solution required $30cm^3$ of 0.2MHCl for its complete neutralization. Calculate percentage purity of the carbonate.

Question 8

50.0 cm^3 sample of sulphuric acid was diluted to 1.00 dm^3. A sample of the diluted sulphuric acid was analysed by titrating with aqueous sodium hydroxide. In the titration, 25.0 cm^3 of 1.00M aqueous sodium hydroxide required 20.0 cm^3 of the diluted sulphuric acid for neutralisation.

a) Give the equation for the full neutralisation of sulphuric acid by sodium hydroxide.

b) Calculate how many moles of sodium hydroxide were used in the titration?

c) Calculate the concentration of the diluted acid.

d) Calculate the concentration of the original concentrated sulphuric acid solution.

Chapter nine

WATER OF CRYSTALLIZATION

INTRODUCTION

Water of crystallization (or **water of hydration)** *are the water molecules that are bonded into a crystalline structure of a compound.* Often written as XH_2Oafter formula of hydrate salt; where X is the whole number.

Hydrated salt (or simply **hydrate**) *is a crystalline compound containing molecules of water of crystallization.*

Anhydrous salt (also known as **anhydrate**) *is the one which is formed after removing water of hydration (commonly by heating) from the hydrated compound.*

RELATIONSHIP BETWEEN MASS OF HYDRATED SALT AND OF ANHYDROUS SALT

Consider hypothetical compound$MX.nH_2O$, where n represents number of molecules of water of crystallization.

Then $MX.nH_2O = \text{Hydrated salt}\text{ , H}$

And MX = Anhydrous salt, A

If the hydrate salt is strongly heated so as to eliminate all water of crystallization according to the following equation:

$$MX.nH_2O \rightarrow MX + nH_2O$$

The mole ratio of hydrated salt to anhydrous salt will be 1:1;

And therefore; $\mathbf{n_H = n_A}$

But $n_H = \frac{\text{Mass of hydrated salt}}{\text{Molar mass of hydrated salt}}$;

And $n_A = \frac{\text{Mass of anhydrous salt}}{\text{Molar mass of anydrous salt}}$

Thus $$\frac{\textbf{Mass of hydrated salt}}{\textbf{Molar mass of hydrated salt}} = \frac{\textbf{Mass of anhydrous salt}}{\textbf{Molar mass of anydrous salt}}$$

And hence (by rearranging the above equation)

$$\frac{\textbf{Mass of hydrated salt}}{\textbf{Mass of anhydrous salt}} = \frac{\textbf{Molar mass of hydrated salt}}{\textbf{Molar mass of anhydrous salt}}$$

Note that: because $n_H = n_A$;

$\textbf{Molarity of hydrated salt} = \textbf{Molarity of anhydrous salt}$(But their mass concentrations are not equal because they have different molar masses).

RELATIONSHIP BETWEEN MASS OF HYDRATED SALT AND OF WATER OF CRYSTALLIZATION

Also from$MX.nH_2O \rightarrow MX + nH_2O$;

Mole ratio of hydrated salt ($MX.nH_2O$) to water (H_2O) is 1: n;

Thus if n_w represents number of moles of water, then;

$$\frac{n_H}{n_w} = \frac{1}{n} \text{or} \frac{\mathbf{n_w}}{\mathbf{n_H}} = \mathbf{n}$$

But $n_H = \frac{\text{Mass of hydrated salt}}{\text{Molar mass of hydrated salt}}$;

And $n_w = \frac{\text{Mass of water}}{\text{Molar mass of water}}$

Then it follows that from $\frac{\mathbf{n_w}}{\mathbf{n_H}} = \mathbf{n}$or $n_w = n \times n_H$;

$$\frac{\textbf{Mass of water}}{\textbf{Molar mass of water}} = \mathbf{n} \times \frac{\textbf{Mass of hydrated salt}}{\textbf{Molar mass of hydrated salt}};$$

And after rearranging the above equation becomes;

$$\frac{\textbf{Molar mass of hydrated salt}}{\mathbf{n} \times \textbf{Molar mass of water}} = \frac{\textbf{Mass of hydrated salt}}{\textbf{Mass of water}}$$

Example 1

A sample of copper (II) sulphate hydrate has a mass of 3.97g. After heating; the$CuSO_4$that remains has a mass of 2.54g. Determine the correct formula and name of the hydrate.

Solution

From; $\frac{\text{Molar mass of hydrated}}{\text{Molar mass of anhydrous}} = \frac{\text{Mass of hydrated}}{\text{Mass of anhydrous}}$;

But it is given that:

Mass of hydrated = 3.97g

Mass of anhydrous = Mass remained after heating = 2.54g

Molar mass of anhydrous = 160g/mol

Molar mass of hydrated($CuSO_4.XH_2O$)

= $64 + 32 + (16 \times 4) + X\big((1 \times 2) + 16\big) = 160 + 18X$

Substituting $\frac{160+18X}{160} = \frac{3.97}{2.54} = 1.563$

Then $160 + 18x = 160 \times 1.563$; $160 + 18x = 250$

$18x = 250 - 160 = 90$; $X = \frac{90}{18} = 5$

Hence the correct formula is$CuSO_4.5H_2O$

Name: Copper(II) sulphate pentahydrate

Example 2

A hydrate is determined to be 45.43% water and 54.57% $CaCl_2$. Find chemical formula for this hydrate.

Solution

Given that (From the definition of mass percentage):

Mass of water in 100g of the hydrate = 45.43g

Mass of anhydrous in 100g of the hydrate = 54.57g

Then using; $\frac{\text{Mass of hydrated}}{\text{Mass of anhydrous}} = \frac{\text{Molar mass of hydrated}}{\text{Molar mass of anhydrous}}$

But molar mass of anhydrous $\mathrm{COCl_2}$

$$= \big(59 + (2 \times 35.5)\big)\mathrm{g/mol} = 130\mathrm{g/mol}$$

Thus molar mass of the hydrate,$\mathrm{CoCl_2.XH_2O}$ will be $130 + 18\mathrm{X}$

Substituting$\frac{100}{54.57} = \frac{130+18\mathrm{X}}{130}$; $130 + 18\mathrm{X} = 238$;

$$18\mathrm{X} = 180; \mathrm{X} = \frac{180}{18} = 6$$

Hence the chemical formula of the hydrate is $\mathrm{CoCl_2.6H_2O}$

Alternative solution

% anhydrous$(\mathrm{CoCl_2})$

$$= \frac{\text{Mass of anydrous }(\mathrm{CoCl_2})\text{ in one mole of the hydrate}}{\text{Molar mass of the hydrate }(\mathrm{CoCl_2.XH_2O})} \times 100\%$$

$$= \frac{\text{Molar mass of anydrous }(\mathrm{CoCl_2})}{\text{Molar mass of the hydrate }(\mathrm{CoCl_2.XH_2O})} \times 100\%$$

$$54.57 = \frac{130}{\mathrm{M_r}\text{ of }\mathrm{CoCl_2.XH_2O}} \times 100\%;$$

$$\mathrm{M_r}\text{ of }\mathrm{CoCl_2.XH_2O} = \frac{130 \times 100}{54.57} = 238\mathrm{g/mol}$$

Thus $130 + 18\mathrm{X} = 238; 18\mathrm{X} = 180; \mathrm{X} = \frac{180}{18} = 6$

Example 3

What mass of water crystallization is present in that mass of hydrated sodium carbonate $(\mathrm{Na_2CO_3.10H_2O})$ which contains 4.6g of sodium?

Solution

Mass of Na in one mole of $Na_2CO_3.10H_2O = 23 \times 2 = 46g$

Molar mass of $Na_2CO_3.10H_2O$

$= \big((23 \times 2) + 12 + (16 \times 3) + 10\,(1 \times 2) + 16)\big)$

$= 46 + 12 + 48 + 180 = 286g/mol$

$$\% \text{ Na} = \frac{\text{Mass of Na in one mole of } Na_2CO_3.10H_2O}{M_r \text{ of } Na_2CO_3.10H_2O} \times 100\%$$

$$= \frac{46g/mol}{286g/mol} \times 100\% = 16.08\%$$

Also using; $\% \text{ Na} = \frac{\text{Mass of Na}}{\text{Total mass of } Na_2CO_3.10H_2O} \times 100\%;$

Where $Na_2CO_3.10H_2O$ has 4.6g of sodium (given)

$16.08 = \frac{4.6g}{\text{Total mass of } Na_2CO_3.10H_2O} \times 100\%;$

From which; $\frac{16.08}{100} \times \text{Total mass} = 4.6g$

$$\text{Total mass} = \frac{4.6 \times 100}{16.08} = 28.6g$$

Therefore total mass of $Na_2CO_3.10H_2O$which contain 4.6g of Na is 28.6g

From;

$$\frac{\text{Molar mass of hydrated salt}}{n \times \text{Molar mass of water}} = \frac{\text{Mass of hydrated salt}}{\text{Mass of water}}$$

$\frac{286g/mol}{180g/mol} = \frac{28.6g}{\text{Mass of water}}; 286 \times \text{Mass of } H_2O = 28.6g \times 180$

$$\text{Mass of water} = \frac{28.6 \times 180}{286} = 18g$$

Hence the mass of water is 18g

Example 4

2.86g of hydrated salt of formula$M_2CO_3.nH_2O$ were heated to constant mass. It was found that 1.06g of solid remained;

(i) Calculated the percentage by mass of water of crystallization in the salt.

(ii) If the formula mass of the hydrated salt is 286, find the atomic mass of M.

Solution

(i) Given mass of hydrated salt = total mass = 2.86g

Mass of water of crystallization

$$= \text{Mass of the hydrate} - \text{Mass of solid remained}$$

$$= 2.86\text{g} - 1.06\text{g} = 1.8\text{g}$$

% of water crystallization

$$= \frac{\text{Mass of water of crystallization}}{\text{Total Mass}} \times 100\%$$

$$= \frac{1.8\text{g}}{2.86\text{g}} \times 100 = 63\%$$

Hence the percentage of water of crystallization is 63%

(ii) From;

$$\frac{\text{Molar mass of hydrated salt}}{\text{n} \times \text{Molar mass of water}} = \frac{\text{Mass of hydrated salt}}{\text{Mass of water}}$$

$$\frac{286}{18\text{n}} = \frac{2.86\text{g}}{1.8\text{g}}$$

$$18\text{n} \times 2.86 = 286 \times 1.8;\ \text{n} = \frac{286 \times 1.8}{18 \times 2.86} = 10$$

Since M_r of$M_2CO_3.nH_2O = 286$ and $n = 10$

Then $(2 \times \text{M}) + (12) + (16 \times 3) + 10\big((1 \times 2) + 16\big) = 286$

$2M + 60 + 180 = 286; 2M = 286 - 240 = 46;\ M = 23$

Hence atomic mas of M is 23

Example 5

A hydrate contains 16.08%sodium, 4.2% carbon, 6.99%hydrogen and 72.73%oxygen by mass. If all hydrogen atoms of the hydrate are in the form of water of crystallization, determine its chemical formula.

Solution

Constituent atoms	Na	C	H	O
Percentage composition	16.08	4.2	6.99	72.73
Mass of each in 100g compound	16.08g	4.2g	6.99g	72.73g
Number of mole of each; using $n = \frac{m}{M_r}$	$\frac{16.08g}{23g/mol}$ $= 0.6991mol$	$\frac{4.2g}{12g/mol}$ $= 0.35mol$	$\frac{6.99g}{1g/mol}$ $= 6.99mol$	$\frac{72.73g}{16g/mol}$ $4.5456mol$
Divide by the smallest to get simpler ratio	$\frac{0.6991mol}{0.35\ mol}$ $= 2$	$\frac{0.35\ mol}{0.35\ mol}$ $= 1$	$\frac{6.99\ mol}{0.35\ mol}$ $= 20$	$\frac{4.45\ mol}{0.35\ mol}$ $= 13$

Thus the empirical formula of the compound is $Na_2CH_{20}O_{13}$

Since all hydrogen are in the form of the water of crystallization and each water molecule has 2 hydrogen atoms, there are ten water of crystallization ($10H_2O$) which contain 10 oxygen atoms leaving 3 oxygen atoms in anhydrate part of the compound.

Hence the chemical formula of the compound is $Na_2CO_3.10H_2O$

Example 6

0.166mol of hydrated salt on strong heating save 17.94g of water. Find the number of moles of water of crystallization in one mole of the hydrated salt?

Solution

Using $\frac{n_w}{n_H} = n$ where n_w and n_H are number of moles of water and the hydrate respectively and n is the number of moles of water of crystallization.

But $n_w = \frac{\text{mass of water}}{\text{molar mass of water}} = \frac{17.94g}{18gmol^{-1}} = \frac{1794}{1800}$ mol

Thus $n = \frac{1794}{1800 \times 0.166} = 6$

Hence the number of moles of water of crystallization is 6

Example 07

A hydrated salt contains 10 moles of water of crystallization per mole of crystal and the percentage by mass of anhydrous salt is 37.06. Find the molar mass of the hydrated salt.

Solution

From $\frac{\text{Mass of hydrated salt}}{\text{Mass of unhydrous salt}} = \frac{\text{Molar mass of hydrated salt}}{\text{Molar mass of unhydrous salt}}$

But mass of anhydrous salt in 100g of hydrated salt is 37.06g

Thus $\frac{\text{Mass of hydrated salt}}{\text{Mass of unhydrous salt}} = \frac{100}{37.06}$

And therefore $\frac{\text{Molar mass of hydrated salt}}{\text{Molar mass of unhydrous salt}} = \frac{100}{37.06}$

If we let molar mass of the hydrate be H, then the molar mass of anhydrous will be $H - 180$ (Mass of 10 moles of water $= 180$)

It follows that $\frac{H}{H-180} = \frac{100}{37.06}$; $37.06H = 100H - 18000$

From which;

$$100H - 37.06H = 62.94H = 18000; H = \frac{18000}{62.94} = 286$$

Hence molar mass of hydrated salt = $286g/mol$.

Example 8

A sample of hydrated calcium sulphate $CaSO_4.XH_2O$ has a relative formula mass of 172. What is the value of X?

Solution

Formula mass of $CaSO_4.XH_2O = 172$

Then $40 + 32 + (16 \times 4) + X((1 \times 2) + 16) = 172$

$40 + 32 + 64 + 18X = 172; 18X = 172 - 136$

$18X = 36; X = \frac{36}{18}; X = 2$

Hence the value of X is 2

Example 9

A teacher told a student that the amount of hydrated salt in lansfordite was 0.03mol and the amount of water lost on heating was 0.15mol. Calculate value of X in the formula $MgCO_3.XH_2O$.

Solution

Using $\frac{n_w}{n_H} = n$ where n_w and n_H are number of moles of water and the hydrate respectively and n is the number of moles of water of crystallization.

Substituting $n = \frac{0.15 \text{ mol}}{0.03 \text{ mol}}; n = 5$

Example 10

A sample of hydrated magnesium sulphate, $MgSO_4.XH_2O$ is found to contain 51.1% water. What is the value of X?

Solution

Given $MgSO_4.XH_2O$ contain 51.1% water
Molar mass of $MgSO_4.XH_2O$

$$= 24 + 32 + (16 \times 4) + X((1 \times 2) + 16)$$

$$= 24 + 32 + 64 + 18X = 120 + 18X$$

Molar mass of anhydrous $MgSO_4 = 120g/mol$

Mass of water in 100g of the hydrate = 51.1g

Mass of the anhydrous $= 100g - 51.1g = 48.9g$

Then using; $\frac{\text{Mass of hydrated}}{\text{Mass of anhydrous}} = \frac{\text{Molar mass of hydrated}}{\text{Molar mass of anhydrous}}$

Substituting $\frac{100}{48.9} = \frac{120+18X}{120}$; $12000 = 5868 + 880.2X$

$880.2X = 6132; X = \frac{6132}{880.2} = 7$

Hence the value of X is 7.

Example 11

A 140g sample of $NiSO_4 .XH_2O$ is heated until no further decrease in mass. The mass of anhydrous salt is 77.5g. Find the number of water molecules in the formula of this hydrate of nickel (II) sulphate and calculate the percentage of water.

Solution

Given the mass of $NiSO_4 .XH_2O = 140.5g$

And the mass anhydrous salt = 77.5g

Molar mass of $NiSO_4 .XH_2O$

$$= 59 + 32 + (4 \times 16) + X((1 \times 2) + 16) = 155 + 18X$$

Molar mass of the anhydrous $= 155g/mol$

From; $\frac{\text{Molar mass of hydrated}}{\text{Molar mass of anhydrous}} = \frac{\text{Mas of hydrated}}{\text{Mass of anhydrous}}$

$$\frac{155+18X}{155} = \frac{140.5}{77.5}\ ; 155 + 18X = 281$$

$$18X = 126 : X = \frac{126}{18} = 7$$

Hence the number of water molecules is 7

Molar mass of the hydrate = $155 + 18X$

= $155 + 18 \times 7 = 281$

Mass of H_2O in one mole of the hydrate $= 18X$

$$= 18 \times 7 = 126$$

$$\%\ H_2O = \frac{\text{Mass of } H_2O \text{ in one mole of the hydrate}}{\text{Molar mass of the hydrate}} \times 100\%$$

$$= \frac{126}{281} \times 100\% = 44.8\%$$

Hence the percentage of $H_2O = 44.8\%$

Example 12

8.58g of washing soda were made up to $250\text{cm}^3$of aqueous solution. $25\text{cm}^3$of this solution required 30cm^3 of 0.2MHCl for its complete neutralisation. Calculate x in the formula, $Na_2CO_3.xH_2O$, for washing soda.

Solution

HCl reacts with washing soda ($Na_2CO_3.xH_2O$) according to the following equation:

$$Na_2CO_3.xH_2O + 2HCl \rightarrow 2NaCl + CO_2 + (x+1)H_2O$$

From which $n_a = 2, n_b = 1$

Then using $\frac{M_aV_a}{M_bV_b} = \frac{n_a}{n_b}$

Where: $M_a = 0.2M, V_a = 30cm^3, V_b = 25cm^3$

Substituting $\frac{0.2\times30}{M_b\times25} = \frac{2}{1}; M_b = \frac{0.2\times30}{2\times25} = 0.12M$

Thus molarity of washing soda is 0.12M

But mass concentration of the washing soda

$$= \frac{\text{Mass}}{\text{Volume}} = \frac{8.58g}{0.25dm^3} = 34.32gdm^{-3}$$

Then from molarity $= \frac{\text{mass concentration}}{\text{molar mass}}$;

$$M_r = \frac{\text{mass concentration}}{\text{molarity}} = \frac{34.32gdm^{-3}}{0.12moldm^{-3}} = 286gmol^{-1}$$

Thus molar mass of washing soda is$286gmol^{-1}$

It follows that:

$$(2 \times 23) + 12 + (16 \times 3) + 18x = 286 \text{ or } x = 10$$

Hence the value of *x* is 10

Example 13

An organic acid has the formula, $H_2C_2O_4.xH_2O$

Calculate *x* from the following data. 1.4301g of the acid was completely neutralized by $22.7cm^3$ of 1M NaOH.

Solution

The acid reacts with NaOH according to the following equation:

$$H_2C_2O_4.xH_2O + 2NaOH \longrightarrow Na_2C_2O_4 + (x+2)H_2O$$

From which $n_a = 1, n_b = 2$

Then using $\frac{\text{number of moles of acid}}{\text{number of moles of base}} = \frac{n_a}{n_b}$;

Where number of moles of base $= MV(M = 1M, V = 22.7cm^3)$

$$= 1\text{M} \times \frac{22.7}{1000}\text{dm}^3 = 0.0227\text{mol}$$

Thence substituting $\frac{\text{number of moles of acid}}{0.0227\text{mol}} = \frac{1}{2}$;

From which number of moles of acid

$$= \frac{1}{2} \times 0.0227\text{mol} = 0.01135\text{mol}$$

Then using $\text{M}_\text{r} = \frac{\text{m}}{\text{n}} \left(\text{From n} = \frac{\text{m}}{\text{M}_\text{r}}\right)$;

Molar mass of the acid = $\frac{1.4301\text{g}}{0.01135\text{mol}} = 126\text{g/mol}$

Then $2 + 24 + 64 + 18\text{x} = 126$ or $\text{x} = 2$

Hence the value of *x* is 2.

PRACTICE EXERCISE 9

Question 1

A calcium chloride hydrate has a mass of 4.72g. After heating for several minutes the mass of hydrate is found to be 3.56g. Use this information to find the formula for the hydrate.

Question 2

When 5g of $FeCl_3.XH_2O$ are heated, 2g of water are driven off.

(i) Find the value of X
(ii) Give the name of the hydrate

Question 3

If the formula of a hydrate is $MgSO_4.XH_2O$ and the percentage by mass of water of crystallization is 51.22. What is the value of X?

Question 4

A 2.815g of $CuSO_4$ was heated until all of the water was removed. Calculated the percentage of water of hydration and the formula of the hydrate if the residue after heating is weighed 2.485g.

Question 5

If 0.715 g of a hydrated form of sodium carbonate exactly reacts with $50cm^3$ of 0.10 M hydrochloric acid, determine the number of moles of water of crystallization present in one mole of the hydrated salt.

Question 6

1.33g of hydrated ethanedioic acid,$H_2C_2O_4.nH_2O$, were dissolved in deionised water and the solution made up to $250cm^3$ in a graduated flask. $25.0cm^3$of this solution were titrated by $21.1cm^3$ of aqueous sodium hydroxide of concentration $0.100moldm^{-3}$. Calculate the number of molecules of water of crystallisation in the hydrated ethanedioic acid.

Chapter ten

FARADAY'S LAWS OF ELECTROYSIS

FARADAY'S FIRST LAW OF ELECTROLYSIS

It states that: *The mass of substance liberated by electrolysis is directly proportional to the quantity of electricity which is passed.*

The quantity of electricity, Q, is measured in coulombs (C) such that: $Q = It$

Where: I is current in amperes (A), t is time in seconds (s)

Thus from the law: $m \propto Q$ or $m \propto It$

Introducing the constant for proportionality, the equation become: $m = ZIt$ or $m = ZQ$

Where Z, is the proportionality constant which is known as **electrochemical equivalent of an element (E.C.E)** whose value depends on the nature of a substance

Definition of electrochemical equivalent of an element

This is the mass of the element liberated by the passage of one coulomb of electricity i.e. one ampere for one second.

Thus $Z = \frac{m}{Q}$ or $Z = \frac{m}{It}$

FARADAY'S SECOND LAW OF ELECTROLYSIS

It states that: *When the same quantity of electricity is passed through different electrolytes, the masses of the different substances liberated are directly proportional to their equivalent weights.*

Where equivalent weight is given by the following formula:

$$\text{Equivalent weight, E} = \frac{\text{Atomic mass}}{\text{Magnitude of ionic charge}}$$

RELATIONSHIP BETWEEN EQUIVALENT WEIGHT AND ELECTROCHEMICAL EQUIVALENT OF AN ELEMENT

Assuming the same amount of electricity Q is passed through two different electrolytes such that:

m_1 is the mass of substance liberated in the first electrolytic cell,

m_2 is the mass of substance liberated in the second electrolytic cell.

Then **from the Faraday's first law of electrolysis:** $m \propto Q$ or $m = ZQ$

Then $m_1 = Z_1Q$ and $m_2 = Z_2Q$

Then $\mathbf{\frac{m_1}{m_2} = \frac{Z_1}{Z_2}}$

(Is the form of the Faraday's first law if there is passage of the same quantity of electricity in the two electrolytes).

But **from the Faraday's second law of electrolysis:**

$m \propto E$ or $m = kE$ where k is the proportionality constant

Then $m_1 = kE_1$ and $m_2 = kE_2$

It follows that: $\mathbf{\frac{m_1}{m_2} = \frac{E_1}{E_2}}$ (Is another form of the Faraday's second law)

Combining Faraday's first and second law by substituting;

$$\frac{E_1}{E_2} \text{ for } \frac{m_1}{m_2} \text{ in } \frac{m_1}{m_2} = \frac{Z_1}{Z_2} \text{ gives } \mathbf{\frac{E_1}{E_2} = \frac{Z_1}{Z_2}}$$

Alternatively the above relationship can be re-written as; $\frac{E_1}{Z_1} = \frac{E_2}{Z_2} = \text{constant}$

That is $\frac{E}{Z}$ is constant and the constant is known as **Faraday's constant, F**

Hence $\mathbf{F} = \frac{\mathbf{E}}{\mathbf{Z}}$

But $Z = \frac{m}{Q}$

Then substituting $\frac{m}{Q}$ for Z in $F = \frac{E}{Z}$ gives $\mathbf{F} = \frac{\mathbf{EQ}}{\mathbf{m}}$

The last equation is the summation of Faraday's first law and Faraday's second law of electrolysis.

Definition of Faradays constant

Faraday constant *is the quantity of electricity that passes when one mole of univalent element (or half a mole of a divalent element, etc.) is discharged or dissolved in electrolysis.* Its value is 96500C i.e. **1F = 96500C.**

Example 1

The electrochemical equivalent of silver is 0.00112gC^{-1}. What mass of silver is deposited by the passage of a steady current of 0.5ampere for 1hour in a silver plating bath?

Solution

Using $m = ZIt = 0.00112 \times 0.5 \times 60 \times 60\text{g} = 2.016\text{g}$

Hence mass of silver deposited is 2.016g

Example 2

0.198g of copper is deposited on a cathode in 40 minutes by passing a steady current of 0.25ampere through copper (II) sulphate solution. Calculate the electrochemical equivalent of copper.

Solution

Using $Z = \frac{m}{Q} = \frac{m}{It}$

Where m = 0.198g, I = 0.25A and t = 40min = 40 × 60s

Substituting $Z = \frac{0.198}{0.25 \times 40 \times 60} gC^{-1} = 3.3 \times 10^{-4} gC^{-1}$

Hence the electrochemical equivalent of copper is $3.3 \times 10^{-4} gC^{-1}$

Example 3

A steady current of 0.27A passed for half an hour in a water voltameter liberated $56cm^3$ of hydrogen at s.t.p.

a) Calculate the electrochemical equivalent of hydrogen

b) What are the electrochemical equivalents of

(i) Oxygen

(ii) Aluminium

Solution

a) At s.t.p:

Mass of hydrogen gas in $22400cm^3$ $(22.4dm^3)$ is 2g

Thus mass of hydrogen gas in$56cm^3 = \frac{56}{22400} \times 2g = 5 \times 10^{-3}g$

Whence mass of hydrogen liberated in electrolysis is $5 \times 10^{-3}g$

Using $Z = \frac{m}{Q} = \frac{m}{It} = \frac{5 \times 10^{-3}}{0.27 \times 0.5 \times 3600} gC^{-1} = 1.0288 \times 10^{-5} gC^{-1}$

Hence electrochemical equivalent of hydrogen is $1.0288 \times 10^{-5} gC^{-1}$.

b) Atomic mass of hydrogen = 1

Magnitude of charge of H^+ = 1

Then equivalent weight of hydrogen, $E_1 = \frac{1}{1} = 1$

Atomic mass of oxygen = 16

Magnitude of charge of O^{2-} = 2

Then equivalent weight of oxygen, $E_2 = \frac{16}{2} = 8$

Atomic mass of aluminium = 27

Magnitude of charge of $Al^{3+} = 3$

Then equivalent weight of aluminium $E_3 = \frac{27}{3} = 9$

But electrochemical equivalent of hydrogen, $Z_1 = 1.0288 \times 10^{-5} gC^{-1}$

And $\frac{E}{Z} = \text{constant}$

Then: (i) $\frac{E_1}{Z_1} = \frac{E_2}{Z_2}$; $\frac{1}{1.0288 \times 10^{-5}} = \frac{8}{Z_2}$; $Z_2 = 8.2304 \times 10^{-5} gC^{-1}$

Hence electrochemical equivalent of oxygen is$8.2304 \times 10^{-5} gC^{-1}$

(ii) $\frac{E_1}{Z_1} = \frac{E_3}{Z_3}$; $\frac{1}{1.0288 \times 10^{-5}} = \frac{9}{Z_3}$; $Z_3 = 9.2592 \times 10^{-5} gC^{-1}$

Hence electrochemical equivalent of aluminium is $9.2592 \times 10^{-5} gC^{-1}$

Example 4

How long must 20A current flow through a solution of $ZnSO_4$ in order to produce 25g of zinc metal?

Solution

Zinc is discharged according to the following equation:

$$Zn^{2+} + 2e \rightarrow Zn$$

From which 2mol of $e^- = 2F = 2 \times 96500C$ liberate 65g (one mole) of Zn.

Let Q be quantity electricity required to produce 25g

So $2 \times 96500C \equiv 65g$

$$Q \equiv 25g$$

Thus (by cross multiplication);

$$Q = \frac{25g \times 2 \times 96500C}{65g} = 7.42 \times 10^4 C$$

Then from $Q = It$;

$$t = \frac{Q}{I} = \frac{7.42 \times 10^4 C}{20A} = 3710s = \frac{3710}{60} \min = 62\min (1\min = 60s)$$

Hence the time taken is $62 \min$

Example 5

How much charge is required to reduce 1mol of Al^{3+} to Al?

Solution

An equation to show reduction of Al^{3+} to Al;

$$Al^{3+} + 3e \rightarrow Al$$

From which;

$$3\text{mol of } e^- \equiv 3F = 3 \times 96500C = 289500C \text{ liberates } 1\text{mole of Al}$$

Hence the amount of charge required is $289500C$.

Example 6

How much electric charge is required to oxidize 1mol of H_2O to O_2.

Solution

Equation to show oxidation of H_2O to O_2;

$$H_2O \rightarrow 2H^+ + O_2 + 2e$$

Thus 2F oxidize 1mol of H_2O to O_2

But $1F = 96500C$

Therefore $2F = 2 \times 96500C = 193000C$

Hence the amount of charge required is $193000C$.

Example 7

A 40A current flowed through a molten iron (III) chloride for 10 hours. Determine the mass of iron and the volume of chlorine gas (measured at STP) that is produced during this time.

Solution

Iron is discharged as per the following equation:

$$Fe^{3+} + 3e \longrightarrow Fe$$

$IF = 1e = 96500C$

$3mol\ of\ e^- \equiv 3F = 3 \times 96500C = 2.895 \times 10^5C$

Hence 2.895×10^5C liberate 56g of Fe

But it is given that:

$I = 40A$

$t = 10hrs = 10 \times 60 \times 60s$

Thus $Q = It = 40A \times 10 \times 60 \times 60s = 1.44 \times 10^6C$

So:

$2.895 \times 10^5C \equiv 56g$

$1.44 \times 10^6C \equiv ?$

Then by cross multiplication;

Mass of Fe liberated = $\frac{1.44\times 10^6 \times 56}{2.895\times 10^5} = 278.5\text{g}$

Hence mass of Fe liberated is 278.5g

Cl_2 is liberated according to the following equation:

$$2Cl^- \rightarrow Cl_2 + 2e$$

$2\text{mol of } e^- \equiv 2F = 2 \times 96500C = 193000C$

Thus 193000C liberate one mole of chlorine gas.

At STP (0°C and 1 atm), 1mol of Cl_2 is equivalent to 22.4dm^3

Thus at STP, 22.4dm^3 is liberated by passing 193000C of electricity.

That is;

$$193000C \equiv 22.4\text{dm}^3$$

$$1.44 \times 10^6C \equiv \ ?$$

Then by cross multiplication, volume of Cl_2 liberated by passing electricity of $1.44 \times 10^6C = \frac{1.44\times 10^6C \times 22.4\text{dm}^3}{193000C} = 167.13\text{dm}^3$

Thus volume of STP produced by passing the given amount of electricity is 167.13dm^3.

Example 8

What current is required to produce 400L of hydrogen gas measured at STP, from electrolysis of water in 1 hour?

Solution

Time taken = 1 hour = $60 \times 60\text{s} = 3600\text{s}$

Hydrogen gas is discharged according to the following equation:

$$2H^+_{(aq)} + 2e^- \longrightarrow H_{2\,(g)}$$

From which 2mol of $e^{-} = 2F = 2 \times 96500C$ liberate $22.4L$ (one mole) of hydrogen gas.

Let Q be quantity electricity required to produce $400L$ of hydrogen gas.

So $2 \times 96500C \equiv 22.4L$

$$Q \equiv 400L$$

Thus (by cross multiplication);

$$Q = \frac{400L \times 2 \times 96500C}{22.4L} = 3.45 \times 10^{6}C$$

Then from $Q = It$;

$$I = \frac{Q}{t} = \frac{3.45 \times 10^{6}C}{60 \times 60s} = 958.3A$$

Hence the current is $958.3A$

Example 9

Exactly 0.4 faraday electric charge is passed through three electrolytic cell in series, first contains$AgNO_3$, second $CuSO_4$ and third $FeCl_3$ solution. How many grams of each metal will be deposited assuming only cathodic reaction in each cell?

Solution

Given $0.4F = 0.4 \times 96500C = 38600C$

For $AgNO_3; Ag^{+} + e \rightarrow Ag_{(s)}$

From which 1mol of $e^{-} = 1F$ liberates 108g (one mole) of Ag.

Thus $1F \equiv 108g$

And $0.4F \equiv m_{Ag}$

Thence (By cross multiplication);

$$m_{Ag} = \frac{0.4F \times 108g}{1F} = 43.2g$$

From Faraday's second law:

Atomic mass of silver (Ag) = 108

Magnitude of charge of $Ag^+ = 1$

Equivalent weight of silver = $E_{Ag} = \frac{108}{1} = 108$

Atomic mass of copper (Cu) = 64

Magnitude of charge of $Cu^{2+} = 2$

Equivalent weight of copper, $E_{Cu} = \frac{64}{2} = 32$

Atomic mass of iron = 56

Magnitude of charge of$Fe^{3+} = 3$

Equivalent weight of iron, $E_{Fe} = \frac{56}{3} = 18.7$

Then from Faraday's second law;

$m \propto E$; $m = kE$; $\frac{m}{E} = k$

Thence $\frac{m_{Ag}}{E_{Ag}} = \frac{m_{Cu}}{E_{Cu}} = \frac{m_{Fe}}{E_{Fe}}$

Then taking $\frac{m_{Ag}}{E_{ag}} = \frac{m_{Cu}}{E_{Cu}}$;

From which (by making$m_{Cu}$the subject)

$m_{Cu} = \frac{m_{Ag} \times E_{Cu}}{E_{Ag}} = \frac{43.2g \times 32}{108} = 12.8g$

Also taking $\frac{m_{Ag}}{E_{Ag}} = \frac{m_{Fe}}{E_{Fe}}$

From which(by making$m_{Fe}$the subject)

$m_{Fe} = \frac{m_{Ag} \times E_{Fe}}{E_{Ag}} = \frac{43.2g \times 56}{108 \times 3} = 7.5g$

Hence:

Mass of silver is 42.2g

Mass of copper is 12.8g

Mass of iron is 7.5g

Example 10

How long has a current of 3A to be applied through a solution of silver nitrate to coat a metal surface of 80cm^2 with 0.005cm thick layer? Given that: Density of silver is10.5g/cm^3.

Solution

Given that:

Current, I = 3A

Area, A = 80cm^2

Thickness, t = 0.005cm

Density, $\rho = \ 10.5\text{g/cm}^3$

Using mass, $\text{m} = \rho \ \times \text{V}$; where $\text{V} = \text{A} \ \times \text{t}$

Hence m = $\rho \ \times \text{A} \ \times \text{t}$

= $10.5\text{g/cm}^3 \ \times \ 80\text{cm}^2 \ \times 0.005\text{cm}$= 4.2g

So mass of silver deposited is 4.2g

Silver is discharged according to the following equation:

$$\text{Ag}^+_{(\text{aq})} + \text{e}^- \ \rightarrow \ \text{Ag}_{(\text{s})}$$

From which 1mol of $\text{e}^- = 1\text{F} = 96500\text{C}$ liberates 108g (one mole) of Ag.

Thus $\qquad 96500\text{C} \ \equiv 108\text{g}$

And $\qquad \text{Q} = \text{It} \ \equiv \ \ 4.2\text{g}$

Thence (by cross multiplication);

$It = 3A \times t = \frac{4.2g \times 96500C}{108g}; t = \frac{4.2 \times 96500}{108 \times 3} s = 1251s$

Hence the time taken is 1251s

Example 11

The following two half – reactions take place in an electrolytic cell with an iron anode and chromium cathode:

$$Fe_{(s)} \rightarrow Fe^{2+}_{(aq)} + 2e^-$$

$$Cr^{3+} + 3e^- \rightarrow Cr_{(s)}$$

During the process, the mass of the iron anode decrease by1.75g. Find the change in mass of chromium cathode.

Solution

It is given that:

$$Fe_{(s)} \rightarrow Fe^{2+}_{(aq)} + 2e^-$$

From which 2mol of $e^- = 2F = 2 \times 96500C$ dissolves 56g (one mole) of Fe.

Thus $\quad 2 \times 96500C \equiv 56g$

And $\quad Q \equiv 1.75g$

Thence (by cross multiplication);

$Q = \frac{1.75g \times 2 \times 96500C}{56g} = 6031.25C$

Hence the time amount of electricity passed was 6031.25C

Also it is given that: $Cr^{3+} + 3e^- \rightarrow Cr_{(s)}$

From which 3mol of $e^- = 3F = 3 \times 96500C$ liberates 52g (one mole) of Cr.

Thus $\quad 3 \times 96500C \equiv 52g$

And $\quad 6031.25 \equiv m_{Cr}$

Thence (by cross multiplication);

$m_{Cr} = \frac{6031.25C \times 52g}{3 \times 96500C} = 1.08g$

Hence the change in mass of in the chromium cathode is 1.08g.

Alternative solution

Using equivalent weight $= \frac{\text{atomic mass}}{\text{magnitude of charge}}$;

$\text{Equivalent weight of Fe}, E_1 = \frac{56}{2} = 28$

$\text{Equivalent weight of Cr}, E_2 = \frac{52}{3}$

Then from Faraday's second law of electrolysis;

$$\frac{m_2}{m_1} = \frac{E_2}{E_1}$$

Substituting $\frac{m_2}{1.75g} = \frac{52}{3 \times 28}$; $m_2 = \frac{52 \times 1.75g}{3 \times 28} = 1.08g$

Hence the change in mass of in the chromium cathode is 1.08g.

Example 12

The chromium in an electrolytic cell increase in mass by 1.37g in 25.5 minutes at a current of 5A. Calculate the charge on the chromium ions in the solution.

Solution

Atomic mass of Cr = 52

Number of moles of $Cr_{(s)}$ produced = $\frac{1.37g}{52g/mol} \left(n = \frac{m}{M_r}\right)$

= 0.0263mol

Quantity of electricity (Q) required to deposit that amount of $Cr_{(s)}$ = It = $5 \times 25.5 \times 60C = 7650C$

But $96500C = IF$

Then by cross multiplication $7650C = \frac{7650 \times IF}{96500} = 0.0793F$

Thus $0.0263mol$ (of Cr) $\equiv 0.0793F$

Then by cross multiplication: 1 mol (of Cr) $\equiv \frac{0.0793F}{0.0263} = 3F$

But number of Faradays required to discharge one mole of the metal is equal to the magnitude of charge of the ion. And because chromium is metal, its charge must be positive.

Hence the charge on chromium ion is +3.

Example 13

In an electrolysis experiment, a current was passed for 5 hours through two cells connected in series. The first cell contains a solution of gold salt and the second cells contain coper (II) sulphate solution. 9.85g of gold was deposited in the first cell. If the oxidation number of gold is one. Find the following:

(a) The amount of copper deposited on the cathode in second cell.
(b) Magnitude of current in ampere.

Solution

Using equivalent weight $= \frac{\text{atomic mass}}{\text{magnitude of charge}}$;

Equivalent weight of Au (gold), $E_1 = \frac{197}{1} = 197$ (it is given that oxidation number of gold is 1 and therefore its magnitude of charge will be 1 too).

Equivalent weight of Cu, $E_2 = \frac{64}{2} = 32$

Then from Faraday's second law of electrolysis;

$$\frac{m_2}{m_1} = \frac{E_2}{E_1}$$

Substituting $\frac{m_2}{9.85g} = \frac{32}{197}$; $m_2 = \frac{32 \times 9.85g}{197} = 1.6g$

Hence the amount of copper deposited is1.6g.

(b) $Au^+ + e \rightarrow Au$

Thus $96500C \equiv 197g$

$$Q \equiv 9.85g$$

$$Q = \frac{96500C \times 9.85g}{197g} = 4825C$$

But $Q = It$; $I = \frac{Q}{t}$; Where time, t = 5 hours = $(5 \times 60 \times 60)s$

Thus $I = \frac{4825C}{5 \times 60 \times 60s} = 0.268A$

Hence the magnitude of current is $0.268A$

Example 14

A 0.2964g of copper was deposited on a passage a current of 0.5A for 30 minutes through a solution of copper (II) sulphate. Calculate the atomic mass of copper.

Solution

Quantity of electricity, Q = It = $0.5 \times 30 \times 60 = 900C$

Copper get deposited according to the following equation:

$$Cu^{2+} + 2e \rightarrow Cu$$

From which $2F = 2 \times 96500C$ liberate one mole of Cu

But one mole of $Cu = M_r =$ Atomic mass

So $2 \times 96500C \equiv M_r$

$$900C \equiv 0.2964g$$

And therefore (by cross multiplication)

$$M_r = \frac{2\times96500\times0.2964}{900} = 63.56$$

Hence the atomic mass of copper is 63.56amu

Example 15

A current of 3.7A is passed for 6 hours between platinum electrode in 0.5L of a 2M solution of $Ni(NO_3)_2$.

(a) What will be the molarity of solution at the end of electrolysis
(b) What will be the molarities of solution if nickel electrode are used.

Solution

(a) Using $Q = It$; where $I = 3.7A$ and $t = 6 \times 60 \times 60s$
Quantity of electricity passed, Q

$$= 3.7A \times 6 \times 60 \times 60s = 79920C$$

Nickel get deposited according to the following equation:

$$Ni^{2+} + 2e \longrightarrow Ni$$

From which $2 \times 96500C \equiv 1mol of\ Ni^{2+}$

Then $79920C \equiv \frac{79920C\times1mol\ of\ Ni^{2+}}{2\times96500C}$

$$= 0.414mol\ Ni^{2+}$$

But in one mole of $Ni(NO_3)_2$, there is one mole of Ni^{2+};

Thus number of moles of $Ni(NO_3)_2$ corresponding to 0.414mol of Ni^{2+} discharged is 0.414mol.

But total number of moles of $Ni(NO_3)_2$ before electrolysis

$$= 2M \times 0.5L = 1mol \quad (n = MV)$$

Thus number of moles of $Ni(NO_3)_2$ remained in the solution

$$= 1mol - 0.414mol = 0.586mol$$

Then using $M = \frac{n}{V}$;

Molarity of $Ni(NO_3)_2 = \frac{0.586\text{mol}}{0.5\text{L}} = 1.172\text{M}$

Hence molarity of the solution is 1.172M

(b) The molarity remains 1.172M because when nickel electrodes are used, anodic nickel will dissolve and get deposited at the cathode and hence the molarity of the solution remain unaffected.

PRACTICE EXERCISE 10

Question 1

Calculate the mass of zinc plated onto the cathode of an electrolytic cell by a current of 750mA in 3.25h.

Question 2

How many grams of gold will be deposited when 5000C pass through a solution by the following equation:

$$AuCl^{-}_{4(aq)} + 3e^{-} \rightarrow Au_{(s)} + 4Cl^{-}_{(aq)}$$

Question 3

Calculate mass of magnesium produced when a current of 60A is passed through a magnesium chloride solution for 4h.

Question 4

The nickel anode in an electrolytic cell decrease in mass by 1.2g in 35.5min. The oxidation half reaction converts nickel atoms into nickel (III) ions. What is the constant current?

Question 5

In the electrolysis of a molten group II metal chloride, 2.5A of constant current is passed through a cell for 1.28 hours. Use the information provided below to determine the identity of group II metal.

Mass of cathode before application of current: 25.720g
Mass of cathode after application of current: 30.949 g

Question 6

A current of 2.68A is passed for one hour through aqueous solution of copper (II) sulphate using copper electrodes. Calculate the change in mass of cathode and that of anode.

EXAMINATION QUESTIONS

Question 1

a) Aluminium sulphate is electrolyzed in the electrolytic cell:
   i) Write an equation to show reduction of aluminium ions at cathode
   ii) What quantity of electricity is needed to deposit 1 mole of aluminium during the electrolysis?

b) 20 cm^3 of sulphuric acid required 25cm^3 of 0.1M potassium hydroxide for complete neutralization. Calculate:
   i) The molarity of the acid
   ii) The concentration, in grammes per litre, of the acid.

Question 2

a) Define the following terms:
   i) Empirical formula
   ii) Molecular formula

b) What is the relation between the two terms in (a) above?

c) An organic compound D has a composition of 52.38% carbon, 12.88% hydrogen and 34.74% oxygen and its molecular weight is 46. Determine its:
   i) Empirical formula
   ii) Molecular formula

Question 3

a) Define the following:
   i) Mass percentage of the solution
   ii) Density of the solution
   iii) Mass concentration of the solution

b) Derive the formula that connects the three terms defined in (a) above.

c) Industrial sulphuric acid is usually labelled as containing 98% of the acid having a density of $1.84 g/cm^3$.
   i) Calculate the molarity of this industrial acid

ii) Calculate the volume of industrial sulphuric acid that will be required to prepare $2dm^3$ of 0.5M sulphuric solution.

Question 4

a) Calcium sulphate is very important inorganic compound. It occurs naturally as the mineral **gypsum** which is composed of calcium sulphate dihydrate.

i) Write chemical formula of **gypsum**

ii) What is the percentage of oxygen in the **gypsum**?

b) A compound of relative molecular mass of 106 was found to be composed of 43.4% sodium, 11.3% carbon and 45.3% oxygen. Determine its:

i) Empirical formula

ii) Molecular formula

Question 5

a) Calculate each of the following:

i) Number of aluminium ions present in solution if 3g of aluminium sulphate are completely dissolved in water.

ii) Number of electrons needed to discharge 32g of copper from copper (II) sulphate.

b) A certain soil requires 80kg of nitrogen (N) per hectare so as to fulfill plant requirements of N. calculate (in kg) the quantity of ammonia sulphate ($(NH_4)_2SO_4$) fertilizer required to meet this demand.

Question 6

a) The quantity of electricity passed during the electrolysis of copper (II) sulphate was 9650 coulombs. Calculate:

i) The number of moles of the metal deposited

ii) The mass of the copper deposited

b) Solid calcium carbonate ($CaCO_3$) reacts with excess nitric acid (HNO_3) liberating carbon dioxide to form soluble calcium nitrate and water. Calculate the amount of nitric acid needed to dissolve 5g of calcium carbonate.

Question 7

a) i) Explain the meaning of Faraday's constant.
   ii) How many coulombs are required to liberate 8g of calcium?

b) In the electrolysis of copper sulphate, the reactions at the electrodes are:
   Cathode: $Cu^{2+} + 2e \rightarrow Cu(s)$
   Anode: $2H_2O \rightarrow O_2 + 4H^+ + 4e$
   What volume of oxygen measured at STP will be liberated at the anode in the time it will take to plate out 5.8g of copper onto the cathode?
   (Electrochemical equivalent of copper $= 0.00033$)

Question 8

a) State Faraday's laws of electrolysis

b) An element X has a relative atomic mass of 88. When a current of 0.5 amperes was passed through fused chloride of X for 32 minutes and 10 seconds, 0.44g of X were deposited at the cathode:
   i) Calculate the number of Faradays needed to liberate one mole of X
   ii) Write the formula of the X ion
   iii) Write the formula of the hydroxide of X

Question 9

a) 8.48g of sodium carbonate crystals were made up to 250cm^3 of solution. 25cm^3 of this solution neutralized 30 cm^3 of 0.2M hydrochloric acid. Calculate the number of moles of water of crystallization in sodium carbonate crystals.

b) 25cm^3 of impure sulphuric acid containing 5.2g/dm^3 reacted with 25cm^3 of sodium hydroxide solution made by dissolving 4g of NaOH in distilled water to make 1litre solution. Calculate the percentage of:

i) Purity of the acid
ii) Impurity of the acid

Question 10

a) 8.5g of sample of iron required just $75cm^3$ of 3M HCl to dissolve it and give a neutral solution. Calculate the percentage purity of the sample of iron.

b) $5cm^3$ of sulphuric acid solution from an automobile battery required $17.48cm^3$ of 1.95M NaOH solution to neutralize the acid. Determine the concentration of the acid in:

i) $moldm^{-3}$
ii) $gramdm^{-3}$

Question 11

a) Define the following:
   i) Electrochemical equivalent
   ii) Chemical equivalent

b) What relationship is there between an electrochemical equivalent of an element and its chemical equivalent?

c) If the Faraday constant is given as $96500C$, calculate the chemical equivalents of:
   i) Hydrogen
   ii) Oxygen
   iii) Copper
   iv) Silver

Question 12

$20cm^3$ of solution containing $7g/dm^3$ of metal hydroxide, XOH, were exactly neutralized by $25cm^3$ of 0.1M hydrochloric acid

a) Write a balanced chemical equation for the neutralization of the metal hydroxide, XOH and hydrochloric acid

b) Calculate the concentration of the metal hydroxide XOH in moles per dm^3

c) i) Calculate the molar mass of XOH
   ii) Identify element X

Question 13

a) How many moles of electrons are required to produce 27g of Al during electrolysis of molten Al_2O_3?
b) One of the method used for preparation of oxygen is by decomposition of hydrogen peroxide in presence of manganese (IV) oxide catalyst:
   i) Write a balanced chemical equation to illustrate the method
   ii) Calculate the volume of oxygen at STP which is theoretically could be obtained from $50cm^3$ of a solution of hydrogen peroxide containing $68g/dm^3$

Question 14

a) ^{204}J, ^{206}K, ^{207}L and ^{A}M are isotopes of element H whose abundancies are $2\%, 24\%, 22\%$ and $X\%$ respectively. Calculate:
   i) The abundancy of X%
   ii) Mass number, A, of isotope M given that the relative atomic mass of element H is 207
b) Ethyl alcohol, which is present in many beverages has a molecular formula C_2H_6O. If 9.2g of ethyl alcohol are available, calculate the:
   i) Number of molecules of ethyl alcohol in 9.2g of ethyl alcohol
   ii) Percentage by weight of oxygen in 9.2g of ethyl alcohol

Question 15

a) An oxide of ion, 4.5g by mass, was completely reduced by heating it in a certain reducing agent and 3.15g of iron was produced. Calculate the empirical formula of the compound.
b) A sample of impure silver of mass 3.45g was used as the anode in an electrolysis purifying process. The cathode was

made up of pure gold of mass 6.45g after the electrolysis, the cathode was found to weigh 9.66g.

i) Calculate the number of coulombs of electricity passed.
ii) What is the percentage purity of the impure silver?

Question 16

a) 0.02 moles of electrons were passed through a solution of sodium hydroxide using platinum electrodes.
   i) Give the names of the gases evolved at each electrode
   ii) Write ionic equations of the reaction taking place at the electrodes
   iii) Calculate the number of moles of each gas produced and the volume which each gas would occupy at S.T.P
b) 289500 coulombs were required to deposit one mole of a metallic element Q from its aqueous salt solution. Calculate the valence of Q.

Question 17

a) Calculate each of the following:
   i) Number of atoms present in 46g of metallic sodium
   ii) Number of molecules in 11.2 litres of carbon dioxide at S.T.P
b) Assume that you are a chemist in a chemical plant that deals with the production of chlorine gas. You want to produce 100 litres of chlorine gas per hour so that you can reach the company's goal of producing 2400 litres every day. What current of electricity will you allow to flow per hour?

Question 18

a) What is the water of crystallization?
b) 5g of $H_2C_2O_4 \bullet xH_2O$ were made up to 250 cm^3 of aqueous solution and 25 cm^3 of this solution required 15.9 cm^3 of 0.5M sodium hydroxide solution to neutralize it. Find the value of x.

Question 19

a) What volume of 0.1M sulphuric acid would be required to neutralize a mixture of 1g of anhydrous sodium carbonate and 1g of sodium hydroxide.

b) 100cm^3 of 0.05M sulphuric acid were placed in a flask and a small quantity of anhydrous sodium carbonate was added. The mixture was boiled to expel carbon dioxide, cooled and the volume restored by addition of distilled water to 100 cm^3 . 25 cm^3 of the solution now required 18 cm^3 of 0.1M sodium hydroxide solution to neutralize it. What was the mass of sodium carbonate added?

Question 20

a) 5g of a mixture of sodium chloride and sodium carbonate (anhydrous) were made up to 250cm^3 of aqueous solution. 25cm^3 of this solution required 40 cm^3 of 0.1M hydrochloric acid for neutralization. What is the percentage by mass of anhydrous sodium carbonate in the mixture?

b) 3.5g of mixture of potassium carbonate and potassium sulphate (both anhydrous) were made up to 250cm^3 of aqueous solution. 25cm^3 of this solution required 24.6cm^3 of 0.110M hydrochloric acid to neutralize. What is the percentage by mass of potassium carbonate in the mixture?

Question 21

a) 10g of calcium carbonate were dissolved in 250 cm^3 of 1M hydrochloric acid, and the solution was then boiled. What volume of 2M potassium hydroxide solution would be required to neutralize the excess of acid?

b) 25 cm^3 of a solution of potassium carbonate required 26.8cm^3 of 0.1M hydrochloric acid for neutralization. If 50 cm^3 of the potassium carbonate solution are mixed with 20 cm^3 of 0.5M hydrochloric acid, how many cubic centimeters of 0.25M sodium hydroxide solution must be added to make the solution neutral?

Question 22

a) On exposure to air, washing soda, $Na_2CO_3 \bullet 10H_2O$, effloresces to leave sodium carbonate monohydrate. Calculate the loss of mass:

   i) Per 7.15g of washing soda

   ii) Per tonne of washing soda if this process of efflorescence is completed

b) Some zinc sulphate crystals were heated to constant mass with the following results:

Mass of crucible	20.00g
Mass of crucible and crystals	25.74g
Mass of crucible and residual	23.22g

From the data, calculate x in the formula, $ZnSO_4 \bullet xH_2O$

Question 23

The compound has the following composition by mass: 68% carbon, 8% hydrogen, and 32% oxygen. If 100mg of this compound is neutralized with exactly 10mL of 0.05MNaOH in reaction with 1:1 stoichiometry; what is the molecular formula of this compound?

Question 24

a) 4.6g of sodium are added slowly to 10g of water in small dish. What is the total mass of the solution left when the reaction is completely?

b) i) What mass of ammonia would be evolved by heating together 100g of ammonium chloride and 70g of calcium hydroxide?

   ii) Which of the reagents is in excess and by what mass?

Question 25

0.2 Faradays of electricity were passed through solutions of:

(i) Copper (II) sulphate (ii) Dilute sulphuric acid

Calculate:

a) The mass of copper liberated in (i)
b) The volume of hydrogen gas evolved at STP in (ii)

Question 26

a) Define the term **mole**
b) If 10^{20} water molecules are removed from 100mg of water, then find the remaining mass of water.
c) Air contains 22.4% oxygen. How many moles of oxygen atoms are there in 1 litre of air at STP?

Question 27

a) Define the following terms:
   i) Mass concentration
   ii) Molarity
b) Derive an equation to show relationship between the two terms defined in (a) above
c) Calculate molarity of pure water, assuming its density is 1g/cm^3.

Question 28

a) Define the following terms:
   i) Normality
   ii) Strength of the solution
b) How many grams of wet NaOH containing 10% water are required to prepare 1 litre of 0.1N solution.
c) A solution is prepared by dissolving 19.6g of sulphuric acid in 2L of water. Calculate:
   i) Molarity and
   ii) Normality of the solution

Question 29

a) What volume of water should be added to 100cm^3 of nitric acid having specific gravity of 1.4 and 70% strength to give 1M solution?

b) In an experiment, $300cm^3$ of 6M nitric acid is required. Only $50cm^3$ of pure nitric acid is available. Pure anhydrous nitric acid (100%) is a colourless liquid with a density of $1522kg/m^3$. Would it possible to carry out the experiment?

Question 30

Calculate each of the following:

a) Mass of Cl atoms in 113.5g of C_2Cl_6
b) Number of oxygen atoms in 45g of $C_{12}H_{22}O_{11}$
c) Number of moles of $CaBr_2$ to obtain 4.5×10^{21} atoms of Br

Question 31

A farm requires 120kg of nitrogen. What is weight of urea fertilizer (80% by weight$CO(NH_2)_2$) needs to be applied to the soil to meet this demand? Show clearly how you obtain your answer.

Question 32

2.5g of X combines with 4g of Y to form a compound with the formula XY_2. If the relative atomic mass of Y is 80. Determine relative atomic mass of X.

Question 33

If 80g of X combines with 1.5×10^{23} atoms of Y to form X_2Y without any of either element remaining; determine gram atomic weight of X.

Question 34

If **m** atoms of X weight 15g and **4m** atoms of element Z whose atomic weight is 30amu weight 45g, determine the atomic weight of X.

Question 35

If 1.181g of an unknown element X reacts with oxygen to form 1.664g of compound, X_2O_3. What is the atomic weight of element X?

Question 36

Calculate the masses of potassium perchlorate ($KClO_4$) and potassium chloride (KCl) produced from $26g$ of potassium chlorate ($KClO_3$) reacting according to the following equation:

$$4KClO_{3(s)} \rightarrow 3KClO_{4(s)} + KCl_{(s)}$$

Question 37

$200mL$ of $0.105M$ silver nitrate is added dropwise to $25mL$ of sodium chloride solution until all chloride precipitates as silver chloride.

(a) What is the molarity of the original sodium chloride solution?
(b) What mass of silver chloride precipitate?

Question 38

Sodium hydroxide solution can be used to neutralize sulphuric acid solution.

(a) Write a full equation for this neutralization reaction.
(b) What mass of sodium hydroxide is required to neutralize completely 5g of sulphuric acid?

Question 39

$25cm^3$ of air were mixed with $50\ cm^3$ of hydrogen and the mixture was exploded. The volume of gas left was $60cm^3$ after cooling and at the same conditions as before. What is the percentage by volume of oxygen in the air?

Question 40

What volume of carbon dioxide measured at STP can be obtained by heating 75g of calcium carbonate to a high temperature?

Question 41

A $60cm^3$ mixture of N_2O and NO is mixed with excess H_2 and the resulting mixture of gases is exploded. The resulting mixture contains $38cm^3$ of N_2. Determine the volume composition of the original mixture.

Question 42

One litre of oxalic acid of density $1.08g/cm^3$ contains 3.24g of oxalic acid. Find its mass percentage.

Question 43

(a) Sulphuric acid solution of density of $1.8g/cm^3$ contains 24.5% acid by weight. What is the molarity of the solution?

(b) A $100cm^3$ of 0.5M calcium nitrate solution is mixed with $200cm^3$ of 1.25M calcium nitrate solution. Calculate the molar concentration of the final solution.

ANSWERS TO PRACTICE EXERCISES

EXERCISE 1

1. (a) +1 (b) +3

2. (a) +3 (b) +6 (c) +7 (d) +1 (e) +5

3. (a) $\mathrm{Ca} = +2; \mathrm{O} = -2; \mathrm{H} = +1$

   (b) $\mathrm{Cu} = +2; \mathrm{N} = +5; \mathrm{O} = -2$

   (c) $\mathrm{N} = +2; \mathrm{O} = -2$

4. $+2$

5. $1-$

EXERCISE 2

1. 6.96amu
2. 126.86amu
3. 179.55amu
4. 10% for ^{113}In and 90% for ^{115}In
5. 28.0854amu
6. Antimony – 121 (57.2%), antimony – 123 (42.8%)
7. 28.09amu
8. 159.14amu
9. 39% for Ir – 191; 61% for Ir – 193

EXERCISE 3

1. (a) 55.19% (b)35% (c)12g (d) 20% ; (Hint: chemical formula of calcium phosphate is $\mathrm{Ca_3(PO_4)_2}$)
2. 36% of Ca and 64% Cl
3. 39.5% C, 7.8% H and 52.7% O
4. 2% H, 32.7% S and 65.3% O
5. 43.4% Na, 11.3% C and 45.3% O
6. (a) 34.3% Na, 17.9% C and 47.8% O (b) 52.2% C, 13%H and 34.8% O (c) 52.9%Al and 47.1% O (d) 44.8% K, 18.4% S and 36.8% O
7. (a) 33.6g (b) 0.397kg (c) 16.3mg (d) 1.37g
8. 0.33kg
9. 480000g/mol

EXERCISE 4

1. (a) 1.53mol (b)3.42×10^{-6}mol (c)0.04mol (d)0.75mol
2. 1.67×10^{22} atoms
3. 3.1×10^{26} atoms
4. (a) 0.54 mol (b)12.46 mol
5. (a) 1.091×10^{24}atoms (b)4.54725×10^{24}atoms
6. (i) 12 g (ii) 1.29×10^{8} g
7. (i) 87g (ii) 0.014 mol (iii)2.99 g
8. (a) 2.5 mol (b)5 mol (c)15 mol

EXERCISE 5

1. 6% $(^{m}/_{m})$
2. (a)0.63g/L, 0.01mol/L, 0.063%(m/m) (b) 90g/L, 0.25mol/L, 0.025%(m/m) (c) 29g/L, 0.5mol/L, 2.9%(m/m).
   Hint: For dilute solution, density of solution is equal to density of water which is$1g/cm^3$).
3. 12% coal tar solution
4. (a) 26% $(^{m}/_{m})H_2SO_4$(b) 16.61% $(^{m}/_{m})$ KCl
5. 0.35kg
6. $0.358mol/dm^3$
7. 450g
8. (a) 12.8mL (b) 60mL
9. 16.7 cm^3
10. 1696cm^3

EXERCISE 6

1. C_6NHO
2. $CH_2O, C_6H_{12}O_6$
3. MF: P_4O_{10} EF: P_2O_5
4. CrO_2
5. $C_{18}H_{34}O_2$
6. CH_4
7. FeO_2
8. SO_3
9. $C_6H_{12}O_6$
10. $NaCrO_4$

EXERCISE 7

1. 40g
2. 0.27 of H_2, 9.73 for Cl_2
3. (a) 0.486g (b) 33.06g
4. 78.57%

5. 70.42% $CaCO_3$ and 29.58% $MgCO_3$
6. 8.1247g
7. (a) 19.2g (b) 16.6g
8. (a)0.868g (b)0.6578g(c)1.11dm^3
9. 20.57g
10. 0.85g

EXERCISE 8

1. 0.323 M
2. 0.411 M
3. 1.902g
4. 0.6752dm^3
5. 0.292g,0.357M
6. 2.32g
7. 37.06%
8. (b) 0.025mol (c) 0.625M (d) 12.5M

EXERCISE 9

1. $CaCl_2 . 2H_2O$
2. (i) X = 6 (ii) Iron (II) chloride hexahydrate
3. X = 7
4. 11.7%
5. 10
6. 2

EXERCISE 10

1. 2.96g
2. 3.402g
3. 107.44g
4. 2.76g
5. Metal with molar mass of 88g/mol which is Sr (Strontium)
6. Change in mass in both cathode and anode is 3.199g

SOLUTIONS TO EXAMINATION QUESTIONS

Question 1

a) i) $Al^{3+} + 3e \rightarrow Al$

ii) From the above equation (in (i)); to deposit one mole of Al, 3 mole of electrons are required.

But 1mole of $e^- \equiv 1F$

Thus 3moles of $e^- \equiv 3F$

But also $1F = 96500$

Thus $3F = 3 \times 96500 = 289500C$

Hence the quantity of electricity required is $289500C$

b) Equation for the reaction between H_2SO_4 and KOH:

$$H_2SO_4 + 2KOH \rightarrow K_2SO_4 + 2H_2O$$

From which mole ratio of H_2SO_4 to KOH is $1:2$

Thus $n_a = 1$ and $n_b = 2$

Using $\frac{M_aV_a}{M_bV_b} = \frac{n_a}{n_b}$

Where:

M_a is unknown molarity of sulphuric acid
V_a is the volume of sulphuric acid $= 20cm^3$
M_b is the molarity of KOH(base) $= 0.1M$
V_b is the volume of KOH (base)$= 25cm^3$

Then substituting $\frac{M_a \times 20}{0.1 \times 25} = \frac{1}{2}$; $M_a = \frac{1 \times 0.1 \times 25}{2 \times 20} M = 0.0625M$

Concentration in $g/L = \text{Molarity} \times \text{Molar mass}$

But Molar mass of H_2SO_4

$$= \left(2 + 32 + (4 \times 16)\right) g/mol = 98\ g/mol$$

Thus concentration of the acid in g/L

$$= 0.0625 \times 98 = 6.125 \text{ g/L}$$

Hence:

i) Molarity of the acid is 0.0625MConcentration of the acid in g/L is 6.125 g/L**Question 2**

ii) Concentration of the acid in g/L is 6.125 g/L**Question 2**

Question 2

a) i) Empirical formula is the chemical formula which gives simplest ratio of number of atoms present in one molecule (or unit) of the compound.

ii) Molecular formula is the chemical formula which gives exact number of atoms present in the one molecule of the compound.

b) Molecular formula is the positive whole number multiple of empirical formula.
That is: Molecular formula $=$ n $\times$ Emprical formula
Where n $=$ 1,2,3,4,5 etc

c)

Constituent atoms	C	H	O
Percentage compositio n	52.38	12.88	34.74
Mass of each in 100g of the compound	52.38g	12.88g	34.74g
Number of moles of each; n $= \frac{m}{m_r}$	$\frac{52.38\text{g}}{12\text{g/mol}}$ $= 4.365$ mol	$\frac{12.88\text{g}}{1\text{g/mol}}$ $= 12.88$ mol	$\frac{34.74\text{g}}{16\text{g/mol}}$ $= 2.17125$ mol
Divide by smallest to	$\frac{4.365 \text{ mol}}{2.17125 \text{ mol}}$	$\frac{12.88 \text{ mol}}{2.17125 \text{ mol}}$	$\frac{2.17125 \text{ mol}}{2.17125 \text{ mol}}$

get simpler ratio	= 2	= 6	= 1

i) The empirical formula is C_2H_6OLet the molecular formula be $(C_2H_6O)_n = C_{2n}H_{6n}O_n$

Let the molecular formula be $(C_2H_6O)_n = C_{2n}H_{6n}O_n$

Then $(2n \times 12) + (1 \times 6n) + 16n = M_r = 46$

$$46n = 46; n = \frac{46}{46} = 1$$

ii) The molecular formula is C_2H_6O**Question 3**

Question 3

a) i) Mass percentage: is the mass of solute in grams dissolved in 100g of solution.

ii) Density of solution: is the ratio of total mass of solution to the total volume of the solution.

iii) Mass concentration: is the mass of solute dissolved in a unit volume of the solution.

b) If $\%\left(\frac{m}{m}\right)$ has been converted to fraction by dividing the percentage by 100, the formula for $\%\left(\frac{m}{m}\right)$ will be as follows:

$$\%\left(\frac{m}{m}\right) = \frac{\text{Mass of solute in g}}{\text{Mass of solution in g}}$$

But $\text{mass of solution} = \text{Density of solution} \times \text{Volume of solution}$

Then $\%\left(\frac{m}{m}\right) = \frac{\text{Mass of solute in g}}{\text{Density of solution} \times \text{Volume of solution}}$

But $\frac{\text{Mass of solute in g}}{\text{Volume of solution}} = \text{Mass concentration}$

$$\%\left(\frac{m}{m}\right) = \frac{\text{Mass concentration}}{\text{Density of solution}}$$

Hence the equation connecting the given three terms is;

$$\%\left(\frac{m}{m}\right) = \frac{\text{Mass concentration}}{\text{Density of solution}}$$

c) From $\%\left(\frac{m}{m}\right) = \frac{\text{Mass concentration}}{\text{Density of solution}}$

$$\text{Mass concentration} = \%\left(\frac{m}{m}\right) \times \text{Density of solution}$$

$$= \frac{98}{100} \times 1.84\text{g/ cm}^3 = 1.8032\text{g/cm}^3$$

Thus mass concentration of the given solution

$$= 1.8032\text{g/cm}^3 = \frac{1.8032\text{g}}{0.001\text{dm}^3} = 1803.2\text{g/dm}^3(1\text{cm}^3 = 0.001\text{dm}^3)$$

Then using molarity $= \frac{\text{Mass concentration in g/dm}^3}{\text{Molar mass in g/mol}}$

Where molar mass of H_2SO_4

$$= \left((2 \times 1) + 32 + (4 \times 16)\right)\text{g/mol} = 98\text{g/mol}$$

Whence the molarity $= \frac{1803.2\text{gdm}^{-3}}{98\text{g/mol}} = 18.4\ \text{moldm}^{-3}$

Hence molarity of the sulphuric acid is 18.4M

(i) Using dilution equation; $M_cV_c = M_dV_d$

Where; $M_c = 18.4\text{M}$; $M_d = 0.5\text{M}$, $V_d = 2\text{dm}^3$

Substituting $18.4 \times V_c = 0.5 \times 2$

From which $V_c = \frac{0.5\times2}{18.4}\text{dm}^3$

Hence the volume of industrial acid required is $0.054\ \text{dm}^3$

Question 4

a) i) $CaSO_4 \bullet 2H_2O$

ii) Molar mass of gypsum

$$= \left(40 + 32 + (4 \times 16) + (2 \times 18)\right)\text{g/mol} = 172\ \text{g/mol}$$

Mass of oxygen atoms in one mole of gypsum

$$= \left((4 \times 16) + (2 \times 16)\right)\text{g} = 96\text{g}$$

$$\%\text{oxygen} = \left(\frac{\text{Mass of oxygen atoms in one mole of gypsum}}{\text{Molar mass of gypsum}}\right) \times 100\%$$

$$= \frac{96\text{g}}{172\text{g}} \times 100\% = 55.8\%$$

Hence the percentage of oxygen in gypsum is 55.8%

b)

Constituent atoms	Na	C	O
Percentage composition	43.4	11.3	45.3
Mass of each in 100g of the compound	43.4g	11.3g	45.3g
Number of moles of each; $n = \frac{m}{m_r}$	$\frac{43.4\text{g}}{23\text{g/mol}} = 1.887\text{ mol}$	$\frac{11.3\text{g}}{12\text{g/mol}} = 0.9417\text{ mol}$	$\frac{45.3\text{g}}{16\text{g/mol}} = 2.83125\text{ mol}$
Divide by smallest to get simpler ratio	$\frac{1.887\text{ mol}}{0.9417\text{ mol}} = 2$	$\frac{0.9417\text{ mol}}{0.9417\text{ mol}} = 1$	$\frac{2.83125\text{ mol}}{0.9417\text{ mol}} = 3$

i) The empirical formula is Na_2CO_3
Let the molecular formula be $(\text{Na}_2\text{CO}_3)_n = \text{Na}_{2n}\text{C}_n\text{O}_{3n}$
Then $(2n \times 23) + (n \times 12) + (3n \times 16) = 106n = M_r = 106; n = 1$

ii) The molecular formula is Na_2CO_3

Question 5

a) i) $M_r(Al_2(SO_4)_3)$
$= \left((2 \times 27) + (3 \times 32) + (12 \times 16)\right)$ g/mol $= 342$ g/mol

Then using $n = \frac{m}{M_r}$

Number of moles of $Al_2(SO_4)_3 = \frac{3g}{342g/mol} = 0.00877$ mol

But in 1mol of$Al_2(SO_4)_3$, there are 2mol of Aluminium ions

So in 0.00877mol of $Al_2(SO_4)_3$, there are 2×0.00877mol $=$ 0.01754mol of aluminium ions

Then from $n = \frac{N}{N_A}$; $N = nN_A$

Therefore number of aluminium ions$= 0.01754 \times 6.02 \times 10^{23}$ ions $= 1.0559 \times 10^{22}$ions

Hence there are 1.0559×10^{22}ions aluminium ions

ii) In copper (II) sulphate, there are Cu^{2+} ions which discharge by gaining electrons according to the following equation: $Cu^{2+} + 2e^- \rightarrow Cu$

That is to discharge 1 mol of copper, 2mol of e^- are required.

And therefore to discharge 32g of copper $\frac{32 \times 2}{64}$ mol $=$ 1mol of electrons will be required.

But 1mol of electrons$= 6.02 \times 10^{23}$ electrons

Hence the number of electrons required is 6.02×10^{23} electrons

b) Molar mass of $(NH_4)_2SO_4$
$= \left((2 \times 18) + 32 + (4 \times 16)\right)$g/mol $= 132$g/mol

Mass of N atoms in one mole of $(NH_4)_2SO_4 = 2 \times 14 = 28$g

Thus 28g(of N) $\equiv$ 132g(of $(NH_4)_2SO_4$)
Then 80 kg (of N) $\equiv \frac{132g \times 80kg}{28g}$ (of $(NH_4)_2SO_4$)
$= 377$ kg

Hence 377kg of ammonium sulphate fertilizer will be required.

Question 6

a) Copper is deposited during electrolysis according to the following equation:

$$Cu^{2+} + 2e^{-} \rightarrow Cu(s)$$

From which 2 moles of electrons ≡ 2F of electricity are required to deposit one mole of copper.

But $1F = 96500C$

Thus $2F = 96500 \times 2C = 193000C$

Whence $193000C \equiv 2$ mol (of Cu)

Then the given $9650C \equiv \frac{2\times9650}{193000}\text{mol} = 0.1\ \text{mol}$

i) Hence 0.1mol of copper will be deposited

Using $n = \frac{m}{M_r}$; $m = nM_r$ where $M_r(Cu) = 64g/mol$

Thus mass of copper $= 0.1 \times 64g = 6.4g$

ii) Hence 6.4g of copper will be deposited

b) The balanced chemical equation to show chemical reaction between carbonate and nitric acid is as follows:

$$CaCO_3 + 2HNO_3 \rightarrow Ca(NO_3)_2 + CO_2 + H_2O$$

From which mole ratio of HNO_3 to $CaCO_3$ is 2: 1

That is $\frac{n_{HNO_3}}{n_{CaCO_3}} = 2$

But from $n = \frac{m}{M_r}$; and $M_r(CaCO_3) = (40 + 12 + 3 \times 16) = 100g/mol$; $n_{CaCO_3} = \frac{5g}{100g/mol} = 0.05mol$

Then $\frac{n_{HNO_3}}{0.05mol} = 2$; $n_{HNO_3} = 2 \times 0.05\ \text{mol} = 0.1\ \text{mol}$

Then using $\mathrm{m = nM_r}$ and $\mathrm{M_r(HNO_3) = (1 + 14 + 3 \times 16)g/mol = 63g/mol}$

Mass of $\mathrm{HNO_3} = 0.1 \times 63\mathrm{g} = 6.3\mathrm{g}$

Hence mass of HNO_3 needed is 6.3g

Note: Your answer must be mass in grams because the given amount in question (amount of$\mathrm{CaCO_3}$) was also mass in grams.

Question 7

a) i) Faraday's constant is the quantity of electricity that passes when one mole of univalent element (or half a mole of a divalent atom) is discharged or dissolved in electrolysis. It is equivalent to one mole of electrons gained or lost during electrolysis and its value is 96500C.
(**Be careful:** The question was asked to **explain** the meaning and not just to give the meaning)
ii) Calcium ions get discharged so as to liberate calcium according to the following equation

$$\mathrm{Ca^{2+} + 2e^- \rightarrow Ca}(s)$$

That is to liberate one mole of calcium, 2 moles of electrons are required.

But 1mol of electrons$\equiv 1\mathrm{F} = 96500\mathrm{C}$
Thus 40 g (of Ca) $\equiv 96500\mathrm{C}$

$$8\mathrm{g(of\ Ca)} \equiv \frac{8 \times 96500}{40} = 19300\mathrm{C}$$

Hence 19300C are required

b) Using $\mathrm{m = ZQ}$; $\mathrm{Q} = \frac{\mathrm{m}}{\mathrm{Z}}$

Where m = Mass (of copper) liberated $= 5.8\mathrm{g}$

Z is the electrochemical equivalent (of copper) $= 0.00033\mathrm{g/C}$

Substituting $Q = \frac{5.8g}{0.00033g/C} = 17576C$ = Amount of electricity that will be responsible for evolving O_2 from H_2O as per the following equation:

$$2H_2O \rightarrow O_2(g) + 4H^+ + 2e$$

From which; to evolve one mole ($22.4dm^3$ at STP) of O_2;

$2F = 2 \times 96500C = 19300C$ of electricity are required

That is $19300C \equiv 22.4dm^3$(of O_2)

Then $174576C \equiv \frac{17576 \times 22.4}{193000} dm^3 = 2.04dm^3$

Hence the volume of oxygen liberated is $2.04dm^3$.

Question 8

a) Faraday's first law: The mass of a substance liberated by electrolysis is directly proportional to the quantity of electricity which is passed
Faraday's second law: When the same quantity of electricity is passed through different electrolytes, the masses of the different substances liberated are directly proportional to their equivalent weights.

b) Number of Faradays $= \frac{Q}{\text{Faraday's constant}} = \frac{It}{96500}$

But;

$I = 0.5A$ and $t = 32 \text{ min } 10 \text{ sec} = 32 \times 60 \text{ sec} + 10 \text{ sec} = 1930 \text{ sec}$

Then number of Faradays $= \frac{0.5 \times 1930}{96500} = 0.01F$

Thus 0.01F of electricity is required to deposit 0.44g of X

Also number of moles of X $= \frac{0.44g}{88g/mol} \left(n = \frac{m}{M_r}\right)$

$= 0.005$ molThen

Then $0.005 \text{ mol} = 0.01F$

Thus $1\text{mol} = \frac{0.01}{0.005}\text{F} = 2\text{F}$

i) Number of Faradays needed$= 2\text{F}$

For monoatomic atom:

Number of Faradays needed per one mole = charge of an ion

Thus the charge on the ion is $+2$ and hence:

ii) The formula of X ion is X^{2+}
iii) The formula of the hydroxide of X is $X(OH)_2$

Question 9

a) Sodium carbonate crystals react with hydrochloric acid according to the following equation:

$$Na_2CO_3 \bullet xH_2O + 2HCl \rightarrow 2NaCl + CO_2 + (x+1)H_2O$$

From which mole ratio of the carbonate (base) to the acid is 1:2 and hence $n_a = 2$ and $n_b = 1$

Then using $\frac{M_aV_a}{M_bV_b} = \frac{n_a}{n_b}$

Substituting $\frac{0.2\times30}{M_b\times25} = \frac{2}{1}$; $M_b = \frac{0.2\times30}{2\times25}\text{M} = 0.12\text{M}$

Using mass concentration (in g/dm^3) $= \frac{\text{Mass of solute in g}}{\text{Volume of solution in dm}^3}$

Mass concentration of sodium carbonate crystals ($Na_2CO_3 \bullet xH_2O$) $= \frac{8.48\text{g}}{0.25\text{dm}^3} = 33.92\ \text{g/dm}^3$

But also molarity $= \frac{\text{Mass comcentration in g/dm}^3}{\text{Molar mass}}$

From which molar mass $= \frac{\text{Mass concentration in g/dm}^3}{\text{Molarity}}$

$$= \frac{33.92\text{g/dm}^3}{0.12\text{mol/dm}^3} = 282.67\ \text{g/mol}$$

Thus;

$M_r(Na_2CO_3 \bullet xH_2O) = (2 \times 23) + 12 + (3 \times 16) + 18x = 282.67$

From which $18x = 282.67 - 106;$ $x = \frac{176.67}{18} = 9.815 \approx 10$

Hence the number of water of crystallization is 10.

b) Molarity of NaOH $= \frac{\text{Mass concentration of NaOH in g/L}}{\text{Molar mass of NaOH}}$

But molar mass of NaOH $= (23 + 16 + 1)\text{g/mol} = 40 \text{ g/mol}$

And mass concentration of NaOH $= 4\text{g/L}$

Thus molarity of NaOH $= \frac{4\text{g/L}}{40 \text{ g/mol}} = 0.1 \text{ mol/L}$

NaOH reacts with H_2SO_4 according to the following equation:

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

From which mole ratio of NaOH to H_2SO_4 is 2:1 and therefore $n_b = 2$ and $n_b = 1$

Then using $\frac{M_aV_a}{M_bV_b} = \frac{n_a}{n_b}$

Substituting $\frac{M_a \times 25}{0.1 \times 25} = \frac{1}{2}$; $M_a = \frac{0.1 \times 25}{2 \times 25} M = 0.05M$

Using mass concentration = molarity × molar mass

Mass concentration of $H_2SO_4 = 0.05 \times$ Molar mass of H_2SO_4

But molar mass of $H_2SO_4 = 98$ g/mol

Thus mass concentration of H_2SO_4

$$= 0.05 \times 98 \text{ g/dm}^3 = 4.9\text{gdm}^{-3}$$

Then % purity $= \left(\frac{\text{Mass concentration of pure sample}}{\text{Mass concentration of impure sample}}\right) \times 100\%$

$$= \frac{4.9\text{gdm}^{-3}}{5.2\text{gdm}^{-3}} \times 100\% = 94.23\%$$

Hence percentage purity of the acid is 94.23%

ii)

$$\text{Percentage impurity} = 100 - \text{Percentage purity}$$
$$= (100 - 94.23)\% = 5.77\%$$

Question 10

a) Iron reacts with HCl according to the following equation:

$$Fe + 2HCl \rightarrow FeCl_2 + H_2$$

From which mole ratio of Fe to HCl is $1:2$

That is $\frac{n_{Fe}}{n_{HCl}} = \frac{1}{2}$

Then using:

$\text{Number of moles} = \text{Molarity} \times \text{Volume of solution (in dm}^3)$

$$n_{HCl} = 3 \times \frac{75}{1000} = 0.225 \text{ moles}$$

Thus $\frac{n_{Fe}}{0.225\text{mol}} = \frac{1}{2}$; $n_{Fe} = \frac{1}{2} \times 0.225\text{mol} = 0.1125\text{mol}$

Then using $m = nM_r$; where $M_r(Fe) = 56\text{g/mol}$

$\text{Mass of iron} = 0.1125 \times 56\text{g} = 6.3\text{g}$

$$\%\text{Purity} = \frac{\text{Mass of pure}}{\text{Mass of impure}} \times 100\% = \frac{6.3\text{g}}{8.5\text{g}} \times 100\%$$
$$= 74.12\%$$

Hence the percentage purity of the iron is 74.12%

b) Sulphuric acid reacts with NaOH according to the following equation:

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

From which $n_b = 2$ and $n_a = 1$

Then using $\frac{M_aV_a}{M_bV_b} = \frac{n_a}{n_b}$

Substituting $\frac{M_a \times 5}{1.95 \times 17.48} = \frac{1}{2}$; $M_a = \frac{1.95 \times 17.48}{2 \times 5} M = 3.4086M$

Also mass concentration $=$ Molarity $\times$ Molar mass

As $M_r(H_2SO_4) = 98g/mol$

Mass concentration of H_2SO_4

$$= 3.4086 \times 98gdm^{-3} = 334gdm^{-3}$$

Hence:

i) Concentration in $moldm^{-3} = 3.4086moldm^{-3}$
ii) Concentration in $gramdm^{-3} = 334gramdm^{-3}$

Question 11

a) i) Is the mass of the element liberated by the passage of one coulomb of electricity.
   ii) Is the ratio of mass of an atom to the magnitude of its ionic charge.

b) If E and Z represent chemical equivalent and electrochemical equivalent respectively;
   Then $\frac{E}{Z} = \text{Constant (Faraday's constant)}$

c)
   i) Atomic mass of H $= 1$Valence number of H $= 1$
   Valence number of H $= 1$
   Chemical equivalent$= \frac{1}{1} = 1$
   ii) Atomic mass of O $= 16$Valence number of O $= 2$
   Valence number of O $= 2$
   Chemical equivalent$= \frac{16}{2} = 8$
   iii) Atomic mass of Cu $= 64$Valence number of Cu (common ionic charge) $= 2$
   Valence number of Cu (common ionic charge) $= 2$
   Chemical equivalent$= \frac{64}{2} = 32$
   iv) Atomic mass of Ag $= 108$Valence number of Ag $= 1$
   Valence number of Ag $= 1$
   Chemical equivalent$= \frac{108}{1} = 108$

Question 12

a) $XOH(aq) + HCl(aq) \longrightarrow XCl(aq) + H_2O(l)$

b) From (a) above: $n_a = n_b = 1$Thus $\frac{M_aV_a}{M_bV_b} = \frac{n_a}{n_b} = 1$; $M_aV_a = M_bV_b$

Thus $\frac{M_aV_a}{M_bV_b} = \frac{n_a}{n_b} = 1$; $M_aV_a = M_bV_b$

Where $M_a = 0.1M,\ V_a = 25cm^3,\ V_b = 20cm^3$

Substituting $0.1 \times 25 = M_b \times 20$; $M_b = \frac{0.1\times25}{20}M = 0.125M$

Hence concentration of XOH in $moldm^{-3}$ is $0.125\ moldm^{-3}$

c) From molarity $= \frac{\text{Mass concentration in } gdm^{-3}}{\text{Molar mass}}$;

$$\text{Molar mass} = \frac{\text{Mass concentration in } gdm^{-3}}{\text{Molarity}}$$

Thus molar mass of XOH $= \frac{7gdm^{-3}}{0.125\ moldm^{-3}} = 56\ gmol^{-1}$

i) Molar mass of XOH is $56\ gmol^{-1}$
Atomic mass of $X + 16 + 1 = 56$
Atomic mass of $X = 56 - 17 = 39$

The element with atomic mass of 39 is potassium.

ii) X is potassium

Question 13

a) $Al^{3+} + 3e \rightarrow Al$

Thus 3 moles of electrons are required to produce 27g (1mol) of Al during the electrolysis

b) i) $2H_2O_2 \xrightarrow{mnO_2} 2H_2O + O_2$
ii) $\text{Mass} = \text{Mass concentration} \times \text{Volume of solution}$

Thus mass of H_2O_2 in $50cm^3$ of the solution

$$= \frac{50}{100} \times 68\text{gdm}^{-3} = 3.4\text{g}$$

Using $n = \frac{m}{M_r}$; where $M_r(H_2O_2) = 34\ \text{gmol}^{-1}$

Number of moles of $H_2O_2 = \frac{3.4\text{g}}{34\text{gmol}^{-1}} = 0.1\ \text{mol}$

But from the chemical equation in (i); mole ratioH_2O_2 to O_2 is $2:1$;

Thus 2mol of H_2O_2 produce 1mol (22.4dm^3 at STP) of oxygen gas

Hence 0.1 mol of H_2O_2 produce $\frac{0.1 \times 22.4\text{dm}^3}{2} = 1.12\text{dm}^3$ of oxygen

Alternative solution:

$$2\ \text{mol (of } H_2O_2) \equiv 1\ \text{mol (of } O_2)$$

$$0.1\ \text{mol (of } H_2O_2) \equiv \frac{0.1}{2}\text{mol} = 0.05\ \text{mol (of } O_2)$$

But at S.T.P; $n = \frac{V}{22.4}$;

$$V = 22.4n = 22.4 \times 0.05\text{dm}^3 = 1.12\text{dm}^3$$

Hence the volume of O_2 produced is 1.12dm^3

Question 14

a)

Thus $2\% + 24\% + 22\% + X\% = 100\%$

From which $X = 52$

i) The abundance $X\%$ is 52%Then using $A_r = \frac{M_1P_1 + M_2P_2 + M_3P_3 + M_4P_4}{100}$

Then using $A_r = \frac{M_1P_1 + M_2P_2 + M_3P_3 + M_4P_4}{100}$

Substituting $207 = \frac{204 \times 2 + 206 \times 24 + 207 \times 22 + A \times 52}{100}$

$$20700 = 9906 + 52A\,; \quad 52A = 20700 - 9906$$

Thus$A = \frac{10794}{52} = 207.6 \approx 208$ (Mass number must be whole number).

ii) Mass number of isotope M is 208

b) Molar mass of ethyl alcohol (C_2H_6O)

$= \big((2 \times 12) + (1 \times 6) + 16\big)\text{g/mol} = 46\text{g/mol}$Using

Using $n = \frac{m}{M_r} = \frac{N}{N_A}$

Then substituting $\frac{9.2}{46} = \frac{N}{6.02 \times 10^{23}}$

$N = \frac{9.2 \times 6.02 \times 10^{23}}{46} = 1.204 \times 10^{23}$ molecules

i) Number of molecules of ethyl alcohol is 1.204×10^{23} molecules.

Since the percentage by weight does not depend on the total weight of the sample;

Percentage weight in 9.2g of ethyl alcohol = Percentage weight of oxygen in 46g (one mol) of ethyl alcohol.

$$= \left(\frac{\text{Mass of oxygen in one mole of } C_2H_6O}{\text{Molar mass of } C_2H_6O}\right) \times 100\%$$

$$= \frac{16}{46} \times 100\% = 34.78\%$$The

ii) The percentage weight of oxygen is 34.78%**Question 15**

Question 15

Total mass of oxide of Fe $=$ Mass of Fe atoms $+$ Mass of O atoms

Thus $4.5\text{g} = 3.15\text{g} +$ Mass of oxygen atoms

$$\text{Mass of oxygen atoms} = 4.5\text{g} - 3.15 = 1.35\text{g}$$

Constituent atoms	Fe	O
Mass of each in 4.5g of the compound	3.15g	1.35g
Number of moles of each; $n = \frac{m}{M_r}$	$\frac{3.15g}{56g/mol}$ $= 0.05625mol$	$\frac{1.35g}{16g/mol}$ $= 0.084375mol$
Divide by smallest number to get simpler ratio	$\frac{0.05625}{0.05625} = 1$	$\frac{0.084375}{0.05625} = 1.5$
Multiply by 2 to get simplest ratio	$1 \times 2 = 2$	$1.5 \times 2 = 3$

Hence the empirical formula of the compound is Fe_2O_3

b) The increase in mass at cathode, is the mass of silver deposited during the electrolysis.

Thus mass of Ag(pure) deposited

$$= (9.66 - 6.45)g = 3.21g$$

Silver get deposited according to the following equation:

$$Ag^+ + e \rightarrow Ag$$

From which 1mol of $e = 1F$ is required to deposit 1mol (108g) of Ag.

That is 108g (of Ag) $\equiv$ 96500C (1F of electricity)

Thus 3.21g (of Ag) $\equiv \frac{3.21 \times 96500}{108} C = 2868C$

i) 2868 coulombs of electricity passed

$$\text{Percentage purity} = \frac{\text{Mass of pure}}{\text{Mass of impure}} \times 100\%$$

$$= \frac{3.21g}{3.45g} \times 100\% = 93\%$$

Hence the percentage purity of the impure silver is 93%

Question 16

a) i) Hydrogen gas at cathode
Oxygen gas at anode

ii) At cathode: $2H^{+}(aq) + 2e \rightarrow H_2(g)$
At anode: $4OH^{-}(aq) \rightarrow 2H_2O(l) + O_2(g) + 4e$

0.02 moles of electrons $\equiv 0.02F$ of electricity

From$2H^{+}(aq) + 2e \rightarrow H_2(g)$; to produce 1mol ($22.4\ dm^3$ at STP) of H_2, 2F of electricity are required

That is $2F \equiv 1\ mol\ (22.4\ dm^3)$

$$0.02F \equiv \frac{0.02 \times 1}{2} mol$$

$$= 0.01\ mol\ \left(or\ \frac{0.02 \times 22.4}{2} dm^3 = 0.224\ dm^3\right)$$

Hence: number of moles of $H_2(g)$ produced is 0.1mol and volume of $H_2(g)$ produced is $0.224\ dm^3$

Also from $4OH^{-}(aq) \rightarrow 2H_2O(l) + O_2(g) + 4e$; to produce 1mol ($22.4\ dm^3$) of O_2, 4F of electricity is required

That is $4F \equiv 1\ mol\ (or\ 22.4\ dm^3)$

$$0.02F = \frac{0.02 \times 1}{4} mol$$

$$= 0.005\ mol\ \left(or\ \frac{0.02 \times 22.4}{4} dm^3 = 0.112 dm^3\right)$$

Hence number of moles of $O_2(g)$ produced is 0.005mol and volume of $O_2(g)$ produced is $0.112dm^3$.

b) $\text{Number of Faradays} = \frac{\text{Amount of charge in coulombs}}{\text{Faraday'sconstant (96500)}}$

$$= \frac{289500}{96500} F = 3F$$ (Required to deposit one mole of Q).

For monoatomic element like metals;

Number of Faradays required to deposit one mole of the element = Valence of the element

Hence the valence is 3

Question 17

a) i) Using $n = \frac{m}{M_r} = \frac{N}{N_A}$: Where $M_r(\text{Na atoms}) = 23\text{g/mol}$

$$\frac{46\text{g}}{23\text{g/mol}} = \frac{N}{6.02 \times 10^{23}\text{atom/mol}}$$

$$N = 2 \times 6.02 \times 10^{23}\text{atoms} = 1.204 \times 10^{24}\text{atoms}$$

Hence the number of atoms present is 1.204×10^{24}atoms.

ii) Using $n = \frac{V}{GMV} = \frac{N}{N_A}$; Where $M_r(CO_2) = 44\text{g/mol}$, GMV at STP $= 22.4\text{L}$

$$\frac{11.2\text{L}}{22.4\text{L/mol}} = \frac{N}{6.02 \times 10^{23}\text{molecules/mol}}$$

$$N = 0.5 \times 6.02 \times 10^{23} \text{ molecules}$$

$$= 3.01 \times 10^{23}\text{molecules}$$

Hence the number of molecules is 3.01×10^{23}molecules.

b) Chlorine is produced electrolytically as per the following equation:

$$2Cl^- \rightarrow Cl_2 + 2e$$

Thus to produce one mole (or 22.4 dm^3 at STP) of Cl_2,

$2F = 2 \times 96500\text{C}$ of electricity are required

That is $22.4\text{L} \equiv 19300\text{C}$

Then $100\text{L} \equiv \frac{193000\text{C} \times 100\text{L}}{22.4\text{L}} = 861607\text{C}$

Whence 861607C of electricity are required to produce 100L of chlorine gas.

But it is required to produce this amount of Cl_2 (100L) in one hour:

That is $t = 1\text{hour} = 3600\text{s}$

Then from $Q = It$; $I = \frac{Q}{t}$

Substituting $I = \frac{861607}{3600}\text{A} = 239\text{A}$

Hence the current of electricity per hour is 239A

Question 18

a) Is the fixed number of molecules of water which are chemically bonded to a salt within a hydrated crystalline.

b) reacts with NaOH according to the following equation:

$$2NaOH + H_2C_2O_4 \bullet xH_2O \rightarrow Na_2C_2O_4 + (x+2)H_2O$$

From which mole ratio of the base, NaOH, to the acid,

$H_2C_2O_4 \bullet xH_2O$ is 2:1 and therefore $n_a = 1, \quad n_b = 2$

Then using $\frac{M_aV_a}{M_bV_b} = \frac{n_a}{n_b} = \frac{1}{2}$; $M_a = \frac{1\times0.5\times15.9}{25\times2}M = 0.159M$

Also from molarity $= \frac{\text{Mass concentration in gdm}^{-3}}{\text{Molar mass}}$;

$$\text{Molar mass} = \frac{\text{Mass concentration in gdm}^{-3}}{\text{Molarity}}$$

But from mass concentration of $H_2C_2O_4 \bullet xH_2O$

$$= \frac{5\text{g}}{0.25\text{dm}^3} = 20\text{g/dm}^3 \ (250\text{cm}^3 = 0.25\text{dm}^3)$$

$$\text{Molar mass} = \frac{20\text{g/dm}^3}{0.159\text{M}} = 126\ \text{g/mol}$$

Thus $M_r(H_2C_2O_4 \bullet xH_2O)$

$$= 2 + (2 \times 12) + (4 \times 16) + 18x = 126$$

$$90 + 18x = 126; 18x = 36; \quad x = \frac{36}{18} = 2$$

Hence the value of x is 2

Question 19

a) Sulphuric acid reacts with anhydrous sodium carbonate according to the following equation:

$$H_2SO_4 + Na_2CO_3 \rightarrow Na_2SO_4 + C0_2 + H_2O$$

From which mole ratio of H_2SO_4 to Na_2CO_3 is 1:1

That is $\frac{n_{H_2SO_4}}{n_{Na_2CO_3}} = 1$ or $n_{H_2SO_4} = n_{Na_2CO_3}$

The sulphuric acid also reacts with sodium hydroxide according to the following equation:

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

From which mole ratio of H_2SO_4 to NaOH is $1:2$

That is $\frac{n_{H_2SO_4}}{n_{NaOH}} = \frac{1}{2}$

Using $n = \frac{m}{M_r}$;

Where $M_r(NaOH) = 40g/mol$; $M_r(Na_2CO_3) = 106\ g/mol$

$n_{Na_2CO_3} = \frac{1}{106}mol$ as$n_{H_2SO_4} = n_{Na_2CO_3}$; Number of moles of H_2SO_4 required to react with Na_2CO_3 was also $\frac{1}{106}mol$

Also $n_{NaOH} = \frac{1}{40}mol$ and as $\frac{n_{H_2SO_4}}{n_{NaOH}} = \frac{1}{2}$; $n_{H_2SO_4}$ required to react with NaOH $= \frac{1}{2} \times \frac{1}{40}mol = \frac{1}{80}mol$

Thus total number of moles of H_2SO_4 required to react with both NaOH and Na_2CO_3

$$= \left(\frac{1}{106} + \frac{1}{80}\right)mol = \frac{93}{4240}mol$$

Then from molarity $= \frac{\text{Number of moles}}{\text{Volume of solution in dm}^3}$

$$\text{Volume of solution in dm}^3 = \frac{\text{Number of moles}}{\text{Molarity}}$$

Thus volume of solution of H_2SO_4 required

$$= \frac{\frac{93}{4240}}{0.1} = \frac{93}{4240 \times 0.1}\,\text{dm}^3 = 0.219\ \text{dm}^3 = 219\text{cm}^3$$

Hence the volume of sulphuric acid is 219cm^3

b) After addition of small amount of Na_2CO_3 (limited reactant), unreacted H_2SO_4 reacts with sodium hydroxide according to the following equation:

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

From which $n_a = 1$ and $n_b = 2$

Using $\frac{M_aV_a}{M_bV_b} = \frac{n_a}{n_b}$;

Where $M_b = 0.1M,\ V_a = 25\text{cm}^3,\ V_b = 18\text{cm}^3$

Substituting $\frac{M_a \times 25}{0.1 \times 18} = \frac{1}{2}$; $M_a = \frac{0.1 \times 18}{2 \times 25}M = 0.036M$

Using $n = MV$;

Number of moles of H_2SO_4 in 100cm^3 of its solution

$$= \frac{100}{1000} \times 0.036\ \text{mol} = 0.0036\ \text{mol}$$

$=$ Number of moles of H_2SO_4 left after the reaction with Na_2CO_3

Again using $n = MV$;

Before the reaction withNa_2CO_3, total number of moles of H_2SO_4

$$= \frac{100}{1000} \times 0.05\ \text{mol} = 0.005\ \text{mol}$$

Thus number of moles of H_2SO_4 reacted with Na_2CO_3

$$= (0.005 - 0.0036)\text{mol} = 0.0014\ \text{mol}$$

H_2SO_4 reacts with Na_2CO_3 according to the following equation:

$$H_2SO_4 + Na_2CO_3 \rightarrow Na_2SO_4 + CO_2 + H_2O$$

From which mole ratio of Na_2CO_3 to H_2SO_4 is 1:1 and therefore

$$n_{H_2SO_4} = n_{Na_2CO_3} = 0.0014\text{ mol}$$

Using $m = nM_r$; where $M_r(Na_2CO_3) = 106g/mol$

Mass of $Na_2CO_3 = 0.0014 \times 106g = 0.1484g$

Hence the mass of anhydrous sodium carbonate added was 0.1484g

Question 20

(a) Only sodium carbonate (base) in the mixture of NaCl and Na_2CO_3 will react with hydrochloric acid according to the following equation:

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

From which $n_a = 2$ and $n_b = 1$

Using $\frac{M_aV_a}{M_bV_b} = \frac{n_a}{n_b}$;

Where $M_a = 0.1M,\ V_a = 40cm^3,\ V_b = 25cm^3$

Substituting $\frac{0.1\times40}{M_b\times25} = \frac{2}{1}$; $M_b = \frac{0.1\times40}{2\times25}M = 0.08M$

Using $n = MV$;

Number of moles of Na_2CO_3 in $250cm^3$ of its solution

$$= \frac{250}{1000} \times 0.08\text{ mol} = 0.02\text{ mol}$$

Then using $m = nM_r$; where $M_r(Na_2CO_3) = 106g/mol$

Mass of Na_2CO_3 in $250cm^3$ of the solution

$$= 0.02 \times 106 = 2.12g$$

Then $\%Na_2CO_3 = \frac{\text{Mass of } Na_2CO_3}{\text{Mass of the mixture}} \times 100\%$

$$= \frac{2.12g}{5g} \times 100\% = 42.4\%$$

Hence the percentage of sodium carbonate is 42.4%

(b) Only potassium carbonate in the mixture of K_2CO_3 and K_2SO_4 will react with hydrochloric acid according to the following equation:

$$K_2CO_3 + 2HCl \rightarrow 2KCl + CO_2 + H_2O$$

From which $n_a = 2$ and $n_b = 1$

Using $\frac{M_aV_a}{M_bV_b} = \frac{n_a}{n_b}$;

Where $M_a = 0.11M$, $V_a = 24.6cm^3$, $V_b = 25cm^3$

Substituting $\frac{0.11\times24.6}{M_b\times25} = \frac{2}{1}$; $M_b = \frac{0.11\times24.6}{2\times25} M = 0.05412M$

Usingn $= MV$;

Number of moles of K_2CO_3 in $250cm^3$ of its solution

$$= \frac{250}{1000} \times 0.05412 \text{ mol} = 0.01353 \text{ mol}$$

Using $m = nM_r$; where $M_r(K_2CO_3) = 138g/mol$

Mass of K_2CO_3 in the $250cm^3 = 0.01353 \times 138g = 1.86714g$

$$\%K_2CO_3 = \frac{\text{Mass of } K_2CO_3}{\text{Mass of the mixture}} \times 100\%$$

$$= \frac{1.86714}{3.5} \times 100\% = 53.3\%$$

Hence the percentage of K_2CO_3 is 53.3%

Question 21

(a) $CaCO_3$ reacts with HCl according to the following equation:

$$CaCO_3 + 2HCl \rightarrow CaCl_2 + H_2O + CO_2$$

From which mole ratio of $CaCO_3$ to HCl is $1:2$ and therefore;

$$\frac{n_{CaCO_3}}{n_{HCl}} = \frac{1}{2} \text{ or } n_{HCl} = 2n_{CaCO_3}$$

Using $n = \frac{m}{M_r}$; $M_r(CaCO_3) = 100g/mol$

Number of moles of $CaCO_3 = \frac{10g}{100g/mol} = 0.1\ mol$

Also using $n = MV$;

Number of moles of $HCl = 1 \times \frac{250}{1000} M = 0.25\ mol$

From the above mole ratio, 0.1 mol of $CaCO_3$ will react with;

$$2 \times 0.1mol = 0.2mol \text{ of HCl}$$

But the given number of moles of $HCl = 0.25\ mol$

Thus $(0.25 - 0.2)\ mol = 0.05\ mol$ of HCl remain unreacted after the reaction between HCl and $CaCO_3$ and this unreacted HCl will react with KOH according to the following equation:

$$HCl + NaOH \rightarrow NaCl + H_2O$$

From which mole ratio of HCl to NaOH is $1:1$ and therefore;

$$n_{NaOH} = n_{HCl} = 0.05\ mol$$

From$M = \frac{n}{V}$; $V = \frac{n}{M}$;

Volume of NaOH required $= \frac{0.05}{2} dm^3 = 0.025dm^3$

(b) Potassium carbonate (base) reacts with hydrochloric acid according to the following equation:

$$K_2CO_3 + 2HCl \rightarrow KCl + CO_2 + H_2O$$

From which mole ratio of K_2CO_3 to HCl is $1:2$ and therefore $n_b = 1$ and $n_a = 2$.

Using $\frac{M_aV_a}{M_bV_b} = \frac{n_a}{n_b}$; $M_a = 0.1M$, $V_a = 26.8cm^3, V_b = 25cm^3$

Substituting $\frac{0.1\times26.8}{M_b\times25} = \frac{2}{1}$; $M_b = \frac{0.1\times26.8}{2\times25}M = 0.0536M$

After mixing the $50cm^3$ of this $0.0536M$ K_2CO_3 and $20cm^3$ of 0.5M HCl;

$n_{K_2CO_3} = \frac{50}{1000} \times 0.0536\text{ mol} = 0.00268\text{ mol}$ $(n = MV)$

$n_{HCl} = \frac{20}{1000} \times 0.5\text{ mol} = 0.01\text{ mol}$ $(n = MV)$

But from the above mole ratio 0.00268 mol of K_2CO_3 needs $2 \times 0.00268 = 0.00536$mol of HCl.

Since the given amount of HCl (0.01mol) excess and $(0.01 - 0.00536)\text{mol} = 0.0046\text{ mol}$ of it remain unreacted after the reaction;

The unreacted HCl will react with NaOH according to the following equation:

$$HCl + NaOH \rightarrow NaCl + H_2O$$

From which mole ratio of HCl to NaCl is 1:1

That is $n_{NaCl} = n_{HCl} = 0.00464\text{ mol}$

Then from $M = \frac{n}{V}$; $V = \frac{n}{M}$;

Volume of NaOH required

$$= \frac{0.00464\text{mol}}{0.25\text{moldm}^{-3}} = 0.01856dm^3 = 18.56cm^3$$

Hence the volume of NaOH required is $18.56cm^3$

Question 22

a) Equation to show efflorescence of $Na_2CO_3 \bullet 10H_2O$ to sodium carbonate monohydrate ($Na_2CO_3 \bullet H_2O$)

$$Na_2CO_3 \bullet 10H_2O \rightarrow Na_2CO_3 \bullet H_2O + 9H_2O$$

Where mole ratio of washing soda to water produced is 1:9 and mass of water produced = loss of mass

Using $n = \frac{m}{M_r}$;

$M_r(Na_2CO_3 \bullet 10H_2O)$
$= \big((2 \times 23) + 12 + (3 \times 16) + (10 \times 18)\big) = 286\text{g/mol}$

i) Number of moles of $Na_2CO_3 \bullet 10H_2O = \frac{7.15}{286}$ mol

Then from the mole ratio;

Number of moles of water produced= $\frac{7.15}{286} \times 9$ mol

Using $m = nM_r$; $Mr(H_2O) = 18\text{g/mol}$

Mass of water produced $= \frac{7.15 \times 9 \times 18}{286}\text{g} = 4.05\text{g} = \text{Mass loss}$

Hence the loss of mass is 4.05g

ii) Number of moles of $Na_2CO_3 \bullet 10H_2O$

$$= \frac{10^6}{286} \text{ mol } (1 \text{ tonne} = 10^6\text{g})$$

Then from the mole ratio;

Number of moles of water produced= $\frac{10^6 \times 9}{286}$ mol

The using $m = nM_r$; where $Mr(H_2O) = 18\text{g/mol}$

Mass of water produced

$$= \frac{10^6 \times 9 \times 18}{286}\text{g} = 566434\text{g} = 566.434\text{kg} = 0.566 \text{ tonne}$$

Hence the loss of mass is 0.566 tonne

Alternative solution:

Since one mole of washing soda lose 9 moles of water:

$$\%\text{Loss of mass} = \frac{\text{Mass of 9 moles water}}{\text{Molar mass of washing soda}} \times 100\%$$

$$= \left(\frac{9 \times 18}{286}\right) \times 100\% = 56.6\%$$

Then using mass loss $= \%\text{Loss} \times \text{total mass of washing soda}$

i) Mass loss $= \frac{56.6}{100} \times 7.15\text{g} = 4.05\text{g}$

ii) Mass loss $= \frac{56.6}{100} \times 1\text{tonne} = 0.566 \text{ tonne}$

b)

Mass of crystals (hydrate)

$$= \text{Mass of crusible and crystals} - \text{Mass of crusible}$$

$$= (25.74 - 20)\text{g} = 5.74\text{g}$$

After heating crystals, water vapour escapes from crucible leaving residue of the anhydrous

Mass of anhydrous

$$= \text{Mass of crucible and residue} - \text{Mass of crucible}$$

$$= (23.22 - 20)\text{g} = 3.22 \text{ g}$$

Molar mass of hydrated salt

$$= 65 + 32 + (4 \times 16) + 18\text{x} = 161 + 18\text{x}$$

Molar mass of anhydrous $= 161$ g/mol

Using; $\frac{\text{Mass of hydrate}}{\text{Mass of anhydrous}} = \frac{\text{Molar mass of hydrated}}{\text{Molar mass of anhydrous}}$

Substituting $\frac{5.74}{3.22} = \frac{161+18\text{x}}{161}$; $518.42 + 57.96\text{x} = 924.14$

$$x = \frac{924.14 - 518.42}{57.96} = 7$$

Hence the value of x is 7

Question 23

Given the compound is exactly neutralized by 10mL of 0.05M NaOH.

$n_{NaOH} = V \times M = 10 \times 10^{-3} \times 0.05 = 5 \times 10^{-4}$mol

Then because mole ratio of the compound to NaOH is 1:1;

Number of moles of compound = number of moles of NaOH

Hence number of moles of the compound is 5×10^{-4}mol

Also it was given that mass of the compound = 100mg

$= 100 \times 10^{-3}$g

Then from $n = \frac{m}{M_r}$; $M_r = \frac{m}{n} = \frac{100 \times 10^{-3}}{5 \times 10^{-4}} = 200\text{g/mol}$

Constituent atom	**C**	**H**	**O**
Percentage composition	60	8	32
Mass of each in 200g of compound	$\frac{60}{100} \times 200\text{g}$ $= 120\text{g}$	$\frac{8}{100} \times 200\text{g}$ $= 16\text{g}$	$\frac{32}{100} \times 200\text{g}$ $= 64\text{g}$
Number of moles of each; $n = \frac{m}{M_r}$	$\frac{120\text{g}}{12\text{gmol}^{-1}}$ $= 10\text{mol}$	$\frac{16\text{g}}{1\text{gmol}^{-1}}$ $= 16\text{mol}$	$\frac{64\text{g}}{16\text{gmol}^{-1}}$ $= 4\text{ mol}$

Therefore the molecular formula is$C_{10}H_{16}O_4$

Question 24

a) Using$n = \frac{m}{Mr}$; $M_r(\text{Na atoms}) = 23g/mol$ and $M_r(H_2O) = 18g/mol$

$$n_{Na} = \frac{4.6g}{23g/mol} = 0.2 \text{ mol}$$
$$n_{Na} = \frac{10g}{18g/mol} = 0.56 \text{ mol}$$

Na reacts with H_2O according to the following equation:

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

From which mole ratio of Na to H_2O is $2:2 = 1:1$

That is $n_{Na}(\text{reacted}) = n_{H_2O}(\text{reacted})$

But the given $n_{H_2O} > n_{Na}$; thus H_2O present in excess while Na is limited reactant

Whence at the end of chemical reaction, all Na will reacts and the produced H_2 being gas will escape from the dish. Only unreacted water and the produced NaOH will remain in the dish

Finding mass of water in the dish:

n_{H_2O} reacted $= n_{Na}$ reacted $= 0.2$ mol

Then n_{H_2O} unreacted $= (0.56 - 0.2)\text{mol} = 0.36 \text{ mol}$

Using $m = nM_r$;

Mass of water which remained unreacted in the dish

$$= 0.36 \times 18g = 6.48g$$

Finding mass of NaOH in the dish:

From the equation: $2Na + 2H_2O \rightarrow 2NaOH + H_2$

Mole ratio of Na to NaOH is $2:2 = 1:1$

That is $n_{Na} = 0.2\ mol$

Using $m = nM_r$; where $M_r(NaOH) = 40g/mol$

Mass of NaOH produced $= 0.2 \times 40g = 8g$

Whence total mass in the dish= mass of H_2O + mass of NaOH

$$= (6.48 + 8)g = 14.48g$$

Hence the total mass of the solution left is $14.48g$.

b) NH_4Cl reacts with $Ca(OH)_2$ according to the following equation:

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

From which mole ratio of NH_4Cl to $Ca(OH)_2$ is $2:1$

That is $\frac{n_{NH_4Cl}}{n_{Ca(OH)_2}} = \frac{2}{1} = 2$; $n_{NH_4Cl} = 2n_{Ca(OH)_2}$

Using $n = \frac{m}{M_r}$; where $M_r(NH_4Cl) = 53.5g/mol$ and $M_r(Ca(OH)_2) = 74g/mol$

$$n_{NH_4Cl} = \frac{100g}{53.5g/mol} = 1.87\ mol$$

$$n_{Ca(OH)_2} = \frac{70g}{74g/mol} = 0.946\ mol$$

$$2 \times n_{Ca(OH)_2} = 2 \times 0.946\ mol = 1.892 > n_{NH_4Cl}(1.87\ mol)$$

Since $n_{NH_4Cl} < 2n_{Ca(OH)_2}$, $Ca(OH)_2$ is excess reactant while NH_4Cl is the limited reactant which will be used to find amount of NH_3 formed.

Mole ratio of NH_4Cl to NH_3 is $2:2 = 1:1$

That is $n_{NH_4Cl} = n_{NH_3} = 1.87\ mol$

Using $m = nM_r$; Where $M_r(NH_3) = 17g/mol$

Mass of $NH_3 = 1.87 \times 17g = 31.79g$

Also from $n_{NH_4Cl} = 2n_{Ca(OH)_2}$; $n_{Ca(OH)_2} = \frac{n_{NH_4Cl}}{2}$

1.87 mol of NH_4 will react with $\frac{1.87}{2}$ mol of $Ca(OH)_2 =$ 0.935 mol

Thus number of moles of unreacted $Ca(OH)_2$

$$= (0.946 - 0.935) \text{ mol} = 0.011 \text{ mol}$$

Then again from $m = nM_r$;

Mass of unreacted $Ca(OH)_2 = 0.011 \times 74g = 0.814g$

Hence:

i) Mass of ammonia evolved is 31.79g
ii) The excess reagent is $Ca(OH)_2$ by mass of 0.814g

Question 25

a) $Cu^{2+} + 2e \rightarrow Cu$

Thus $2F \equiv 1$ mol (64g of Cu)

Then $0.2F \equiv \frac{0.2 \times 64}{2} g = 6.4$

Hence the mass of copper liberated in (i) is 6.4g

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

Thus $2F \equiv 1$ mol (22.4 dm^3 of H_2 at STP)

Then $0.2F \equiv \frac{0.2 \times 22.4 \text{ dm}^3}{2} g = 2.24 \text{ dm}^3$

Hence the volume of hydrogen gas evolved is 2.24 dm^3

Question 26

a) Is the amount of substance which contain as many elementary entities as there in 12g of carbon-12 isotope.
b) Finding mass of water removed by using;

$$n = \frac{N}{N_A} = \frac{m}{M_r};$$

Where $N = 10^{20}$ molecules $M_r(H_2O) = 18g/mol$

That is $\frac{10^{20}}{6.02\times10^{23}} = \frac{m}{18}$;

$$m = \frac{10^{19}\times 18}{6.02\times10^{23}} = 0.00299g = 2.99mg$$

Thus mass of water removed $= 2.99$ mg

And mass of water remained

$$= 100mg - 2.99mg = 97.01mg$$

c) Volume of oxygen $= \%$oxygen $\times$ Volume of air

$$= \frac{22.4}{100}\times 1L = 0.224L$$

Using $n = \frac{V}{GMV}$ where at STP GMV $= 22.4L/mol$

Number of moles of oxygen $= \frac{0.224L}{22.4L/mol} = 0.01$ mol of oxygen

But oxygen exists as diatomic molecule (O_2); that is each oxygen molecule has two oxygen atoms.

Whence number of moles of oxygen atoms

$$= 2 \times \text{number of moles of oxygen molecules}$$

$$= 2\times 0.01mol = 0.02mol$$

Hence number of moles of oxygen atoms is 0.02mol.

Question 27

a) i) Is the mass of solute dissolved in a unit volume of solution.
ii) Is the number of moles of solute dissolved in a litre of solution.

b) From molarity $= \frac{\text{Number of moles of solute}}{\text{Volume of solution}}$

But number of moles of solute $= \frac{\text{Mass of solute}}{\text{Molar mass of solute}}$

Then molarity $= \frac{\text{Mass of solute}}{\text{Volume of solution in dm}^3 \times \text{Molar mass of solute}}$

But $\frac{\text{Mass of solute}}{\text{Volume of solution in dm}^3} = \text{Mass concentration in g/dm}^3$

Hence molarity $= \frac{\text{Mass concentration in g/dm}^3}{\text{Molar mass of solute in g/mol}}$

c) Since $1\ \text{cm}^3 = 0.001\text{dm}^3$; $\frac{1\text{g}}{\text{cm}^3} = \frac{1\text{g}}{0.001\text{dm}^3} = 1000\ \text{g/dm}^3$

Thus density of water in g/dm^3 is 1000g/dm^3

But for pure substance (water); Mass concentration $=$ Density

Thus mass concentration of pure water in $\text{g/dm}^3 = 1000\text{g/dm}^3$

The using molarity $= \frac{\text{Mass concentration in g/dm}^3}{\text{Molar mass of solute}}$

Where molar mass of water $= 18\text{g/mol}$

Molarity of pure water $= \frac{1000\text{g/dm}^3}{18\text{g/mol}} = 55.56\ \text{moldm}^{-3}$

Hence the molarity of pure water is 55.56M

Question 28

a) i) Is the number equivalents of a solute dissolved in a litre of solution.

ii) Is the amount of solute dissolved in a given amount of solution.

b) For base (NaOH); Normality $=$ Molarity $\times$ AcidityBut acidity of NaOH is 1

But acidity of NaOH is 1

So Normality $=$ Molarity

Thus molarity of NaOH $= 0.1\text{M}$

From $n = MV$;

Number of moles of NaOH in 1L of solution

$$= 0.1 \times 1\text{mol} = 0.1\text{mol}$$

Then using $m = nM_r$; where $M_r(\text{NaOH}) = 40\text{g/mol}$

Mass of NaOH required $= 0.1 \times 40\text{g} = 4\text{g}$

Also $\%\left(\frac{m}{m}\right)\text{NaOH} + \%\left(\frac{m}{m}\right)H_2O = 100\%$

From which $\%\left(\frac{m}{m}\right)\text{NaOH} = 100\% - \%\left(\frac{m}{m}\right)H_2O$

$$= (100 - 10)\% = 90\%$$

And $\%\left(\frac{m}{m}\right)\text{NaOH} = \frac{\text{Mass of NaOH}}{\text{Mass of solution}} \times 100\%$

Substituting $90 = \frac{4}{\text{Mass of solution}} \times 100$;

$$\text{Mass of solution} = \frac{4 \times 100}{90}\text{g} = 4.44\text{g}$$

Hence 4.44g of wet NaOH are required

c) Mass concentration of the solution

$$= \frac{\text{Mass of sulphuric acid (solute)}}{\text{Volume of solution}} = \frac{19.6\text{g}}{2\text{L}} = 9.8\text{g/L}$$

i) Using $\text{Molarity} = \frac{\text{Mass concentration in g/L}}{\text{Molar mass in g/mol}}$

Where molar mass of $H_2SO_4 = 98\text{g/mol}$

$$\text{Molarity} = \frac{9.8\text{g/L}}{98\text{g/mol}} = 0.1\ \text{mol/L}$$

Hence molarity of the solution is 0.1M

ii) For an acid (H_2SO_4) ;

$$\text{Normality} = \text{Molarity} \times \text{Basicity}$$

But basicity of $H_2SO_4 = 2$

Thus normality of the solution is 0.2N

Question 29

a) Specific gravity of the solution

$$= \frac{\text{Density of solution (nitric acid solution)}}{\text{Density of water } (1\text{gcm}^{-3})}$$

Thus density of nitric acid solution

$= \text{Specific gravity} \times \text{Density of water}$

$$= 1.4 \times 1\text{gcm}^{-3} = 1.4\text{gcm}^{-3} = \frac{1.4\text{g}}{0.001\text{dm}^3} = 1400\text{g/dm}^3$$

Using Mass concentration = Density × Mass percentage

Mass concentration of nitric acid solution

$$= \frac{1400\text{g}}{\text{dm}^3} \times \frac{70}{100} = 980\text{g/dm}^3$$

Also using molarity $= \frac{\text{Mass concentration in g/dm}^3}{\text{Molar mass}}$

Where molar mass of HNO_3

$$= \big(1 + 14 + (3 \times 16)\big)\ \text{g/mol} = 63\text{g/mol}$$

Molarity of HNO_3(before addition of 100cm^3 of water)

$$= \frac{980\text{g/dm}^3}{63\text{g/mol}} = \frac{980}{63}\text{M}$$

Then using dilution formula: $M_cV_c = M_dV_d$

Where $M_c = \frac{980}{63}$; $V_c = 100\text{cm}^3$; $M_d = 1\text{M}$

Substituting $\frac{980}{63} \times 100 = 1 \times V_d$; $V_d = 1555.56\ \text{cm}^3$

But $V_d = V_c + V_{H_2O}$; $V_{H_2O} = V_d - V_c$

$$= (1555.56 - 100)\text{cm}^3 = 1455.56\ \text{cm}^3$$

Hence the volume of water to be added is $1455.56\ cm^3$.

b) Density of pure nitric acid

$$= 1522 kg/m^3 = \frac{1522\ \times 10^3 g}{10^3 dm^3} = 1522 g/dm^3$$

But for pure substance (nitric acid):

$$\text{Density} = \text{Mass concentration}$$

Thus mass concentration of pure nitric acid is also $1522 g/dm^3$

Using molarity $= \frac{\text{Mass concentration in g/dm}^3}{\text{Molar mass}}$

Where molar mass of $HNO_3 = 63 g/mol$

Molarity of pure $HNO_3 = \frac{1522 g/dm^3}{63 g/mol} = \frac{1522}{63} M$

Then it is required to dilute this $\frac{1522}{63} M\ HNO_3$ to $6M\ HNO_3$ and for the experiment to be carried out the volume of $6M\ HNO_3$ must be at least (greater or equal to) $300 cm^3$.

Checking whether the volume of diluted solution ($6M\ HNO_3$) will be enough or not by using dilution formula:

$$M_c V_c = M_d V_d$$

Where $M_c = \frac{1522}{63} M;\ V_c = 50 cm^3;\ M_d = 6M$

Substituting $\frac{1522}{63} \times 50 = 6 \times V_d$;

$$V_d = \frac{1522 \times 50}{63 \times 6} cm^3 = 201.3\ cm^3$$

Thus with the given volume of pure nitric acid, only $201.3\ cm^3$ of $6M\ HNO_3$ may be prepared and not the required least amount of $300 cm^3$ and hence it is not possible to carry out the experiment.

Alternative solution:

Alternatively we may find the volume of pure nitric acid required to prepare 300cm^3 of 6M HNO_3 and checking whether the given volume of pure nitric acid is enough or not to prepare the diluted solution as follows:

Again from dilution formula: $M_cV_c = M_dV_d$

Where $M_c = \frac{1522}{63}\text{M}$; $V_d = 300\text{cm}^3$; $M_d = 6\text{M}$

Substituting $\frac{1522}{63} \times V_c = 6 \times 300\text{cm}^3$;

$$V_c = \frac{6 \times 300 \times 63}{1522}\text{cm}^3 = 74.5\text{ cm}^3$$

Thus 74.5 cm^3 of the pure acid is required to prepare 300cm^3 of 6M HNO_3

But the given volume of pure nitric acid is only 50cm^3(not enough to prepare 300cm^3 of 6M HNO_3) and hence it is not possible to carry out the experiment.

Question 30

a) Molar mass of C_2Cl_6

$$= \big((2 \times 12) + (6 \times 35.5)\big)\text{g/mol} = 237\text{g/mol}$$

Mass of Cl atoms in one mole of $C_2Cl_6 = 6 \times 35.5\text{g} = 213\text{g}$

Thus $237\text{g}(\text{of } C_2Cl_6) \equiv 213\text{g (of Cl atoms)}$

Then$113.5\text{g}(\text{of } C_2Cl_6) \equiv \frac{113.5\times213}{237}\text{g} = 102\text{g (of Cl atoms)}$

Hence the mass of Cl atoms is 102g

b) Molar mass of $C_{12}H_{22}O_{11}$

$$= \big((12 \times 12) + (22 \times 1) + (11 \times 16)\big) = 342\text{g/mol}$$

Number of O atoms in one mole of $C_{12}H_{22}O_{11}$

$$= 11 \times 6.02 \times 10^{23}\text{ atoms}$$

Thus $342g(\text{of } C_{12}H_{22}O_{11}) \equiv 11 \times 6.02 \times 10^{23}$ atoms

Then $45g(\text{of } C_{12}H_{22}O_{11}) \equiv \frac{45 \times 11 \times 6.02 \times 10^{23}}{342}$ atoms

$$= 8.71 \times 10^{23} \text{atoms}$$

Hence the number of oxygen atoms is 8.71×10^{23}atoms

c) In each mole of $CaBr_2$, there are two moles of Br atoms

But 2 moles of Br atoms $= 2 \times 6.02 \times 10^{23}$ atoms

Thus $2 \times 6.02 \times 10^{23}$ atoms (of Br) $\equiv 1$ mole (of $CaBr_2$)

$$4.5 \times 10^{21} \text{ atoms (of Br)} \equiv \frac{4.5 \times 10^{21}}{2 \times 6.02 \times 10^{23}} \text{mol (of } CaBr_2)$$

$$= 0.00374 \text{ mol}$$

Hence 0.00374 mol of $CaBr_2$ are required

Question 31

Molar mass of $(CO(NH_2)_2)$ is 60g/mol = 0.06 kg/mol

Mass of nitrogen in one mole of urea = 28g =0.028kg

Thus, 0.028kg of nitrogen is produced by 0.06kg of urea

And 120kg of nitrogen will be produced by

$$\frac{120 \times 0.06}{0.028} = 257.143\text{kg of urea}$$

But the fertilizer is 80% urea by weight

Therefore 80 kg is contained in 100kg of fertilizer

And 257.143kg of urea will be contained in $\frac{257.143 \times 100}{80}$ kg or 321kg of fertilizer

Hence 321kg of urea fertilizer required to meet the requirement.

Question 32

$$X + 2Y \rightarrow XY_2$$

Number of moles of Y $= \frac{\text{mass of Y}}{\text{Molar mass of Y}} = \frac{4\text{g}}{80\text{g/mol}} = 0.05 \text{ mol}$

But;

$$1 \text{ mol of X} = 2 \text{ mol of Y}$$

$$? \quad = \quad 0.05\text{mol}$$

By crossing multiplication;

Number of moles of X $= \frac{0.05}{2} = 0.025$

But mass of X = 2.5g;

Then from, $n = \frac{m}{M_r}$; $M_r = \frac{m}{n} = \frac{2.5 \text{ g}}{0.025 \text{ mol}} = 100\text{g/mol}$

Hence the relative atomic mass of X is 100

Question 33

Number of moles of atoms of Y, $n = \frac{N}{N_A}$

$$= \frac{1.5 \times 10^{23} \text{ atom}}{6.023 \times 10^{23} \text{ atom/mol}} = 0.25 \text{ mol}$$

Also

$$1 \text{ mol of Y} \equiv 2\text{mol of X}$$

$$0.25\text{mol of Y} \equiv 0.25 \times 2\text{mol} = 0.5\text{mol of X}$$

Also from $n = \frac{m}{M_r}$; $M_r = \frac{m}{n}$;

$$M_r \text{of X} = \frac{80\text{g}}{0.3 \text{ mol}} = 160\text{g/mol}$$

Hence atomic weight of X is 160g/mol

Question 34

Molar mass of Z atoms = 30g/mol

Mass of one atom of Z = $\frac{\text{Molar mass}}{\text{Avogdro's number}}$

$$= \frac{30\text{g/mol}}{6.02\times10^{23}\text{atom/mol}} = 4.98\times10^{-23}\text{g/atom}$$

Then mass of **4m** atoms = Mass of one atom $\times$ **4m** atoms

$$= 4.98\times10^{-23}\text{g/atom}\times\mathbf{4m}\ \text{atom}$$

$$= 1.992\ \times\ 10^{-22}\mathbf{m}\text{g}$$

But it is given mass of 4m atoms = 45g

Thus $1.992\ \times\ 10^{-22}\mathbf{m} = 45;\ \mathbf{m} = \frac{45}{1.992\times10^{-22}} = 2.26\times10^{23}$

Also it is given that **m** atoms of X weighs 15g;

Thus 2.26×10^{23}atoms (value of **m**) weighs 15g.

Thus mass of one atom of X

$$= \frac{15\text{g}}{2.26\times10^{23}\text{atoms}} = 6.64\times10^{-23}\text{g/atom}$$

Then molar mass of X = Mass of one atom of X $\times$ Avogadro's number

$=6.64\times10^{-23}\text{g/atom}\ \times6.02\times10^{23}\text{atom/mol} = 40\text{g/mol}$

Hence the atomic weight of X is 40amu.

Question 35

Mass of O in X_2O_3

$$= \text{Total mas of}\ X_2O_3 - \text{Mass of X in the sample}$$

$$= (1.664 - 1.181)\text{g} = 0.483\text{g}$$

Number of moles of oxygen atoms= $\frac{m_O}{M_O} = \frac{0.483\text{g}}{16\text{gmol}^{-1}} = 0.03\text{mol}$

But in one mole of X_2O_3; 3 mol of O atoms $\equiv$ 2 mol of X atom

$$0.03 \text{ mol of O atoms} \equiv \quad ?$$

By crossing multiplication, number of mole of X corresponding to 0.03mol of oxygen atoms = $\frac{0.03 \times 2}{3}$ mol $= 0.02$mol

Thus number of moles of X atoms corresponding to 1.181g of its mass is 0.02mol

Then using $n = \frac{m}{M_r}$; From which $M_r = \frac{m}{n}$

Molar mass of X atoms = $\frac{1.181g}{0.02mol} = 59.05g/mol$

But numerical value of molar mass of atoms in g/mol is equal to the numerical value of atomic weight in amu.

Hence the atomic weight of element X is 59.05amu.

Question 36

Molar mass of $KClO_3 = 39 + 35.5 + (16 \times 3) = 122.5g/mol$

Number of mole of $KClO_3 = \frac{m}{M_r} = \frac{26g}{122.5g/mol} = 0.2122mol$

From the given reaction equation, mole ratio of $KClO_4$ to $KClO_3$ is 3:4.

That is

$$\frac{n_{KClO_4}}{n_{KClO_3}} = \frac{3}{4};\ n_{KClO_4} = \frac{3}{4} \times n_{KClO_3} = \frac{3}{4} \times 0.2122 = 0.15915mol$$

But molar mass of $KClO_4 = 39 + 35.5 + (16 \times 4) = 138.5g/mol$

Then $m = n \times M_r = 0.15915mol \times 138.5g/mol = 22g$

Hence mass of $KClO_4$ produced is 22g

Also from the same given equation, mole ratio of KCl to $KClO_3$ is 1:4.

That is

$$\frac{n_{KCl}}{n_{KClO_3}} = \frac{1}{4};\ n_{KCl} = \frac{1}{4} \times n_{KClO_3} = \frac{1}{4} \times 0.2122 = 0.05305\text{mol}$$

But molar mass of KCl $= 39 + 35.5 = 74.5\text{g/mol}$

Then $\text{m} = \text{n} \times \text{M}_\text{r} = 0.05305\text{mol} \times 74.5\text{g/mol} = 3.95\text{g}$

Hence mass of KCl produced is 3.95g

Question 37

Number of moles of $AgNO_3$

$$= \text{M} \times \text{V} = 0.105\text{mol/L} \times 200 \times 10^{-3}\text{L} = 0.021\text{mol}$$

From the reaction equation;

$$AgNO_3 + NaCl \rightarrow AgCl + NaNO_3$$

Mole ratio of $AgNO_3$ to NaCl is 1: 1

Thus number of moles of NaCl was also 0.021mol

And it is given that volume of NaCl $= 25\text{mL} = 25 \times 10^{-3}\text{L}$

Then Molarity of NaCl $= \frac{\text{n}}{\text{V}} = \frac{0.021\text{mol}}{25\times10^{-3}\text{L}} = 0.84\text{M}$

Hence molarity of NaCl is 0.84M

(b) Also from the reaction equation:

$$AgNO_3 + NaCl \rightarrow AgCl + NaNO_3$$

Mole ratio of $AgNO_3$ to AgCl is 1: 1

Thus number of moles of AgCl was also 0.021mol

Then using $\text{m} = \text{n} \times \text{M}_\text{r}$; where M_r of AgCl $= 143.5\text{g/mol}$

Mass of AgCl $= 0.021\text{ mol} \times 143.5\text{g/mol} = 3.01\text{g}$

Hence mass of AgCl produced is 3.01g

Question 38

(a) $2NaOH + H_2SO_4 \rightarrow Na_2SO_4 + 2H_2O$

(b) From $n = \frac{m}{M_r}$; where M_r of $H_2SO_4 = 98g/mol$ and given mass of $H_2SO_4, m = 5g$

Thus $n = \frac{5g}{98g/mol} = 0.051mol$

Number of moles of $H_2SO_4 = 0.051mol$

But from the reaction equation;

$H_2SO_4 + 2NaOH \rightarrow Na_2SO_4 + H_2O; n_a = 1, n_b = 2$

Then using $\frac{\text{number of moles of acid}}{\text{number of moles of base}} = \frac{n_a}{n_b}$;

$\frac{0.051mol}{\text{Number of moles of base(NaOH)}} = \frac{1}{2}$;

Number of moles of base $= 2 \times 0.051mol = 1.2mol = 0.102mol$

Then using $m = n \times M_r$; where $M_r$of NaOH $= 40g/mol$

$$\text{Mass of NaOH} = 40g/mol \times 0.102mol = 4.08g$$

Hence mass of NaOH required is 4.08g

Question 39

Only oxygen in the air will react with hydrogen according to the equation:

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$$

From which mole ratio of H_2 to O_2 is $2:1$

That is $\frac{n_{H_2}}{n_{O_2}} = \frac{2}{1} = 2$

Let the volume of oxygen in the air be xcm^3

Then according to Avogadro's law; because volume varies directly proportional to the number of moles;

$$\text{Volume ratio} = \text{mole ratio}$$

Thus $\frac{V_{H_2}}{V_{O_2}} = 2$

That is $\frac{V_{H_2}}{x} = 2; V_{H_2} = 2x$

Whence xcm^3 of O_2 reacts with $2xcm^3$ of H_2

In order to know percentage of oxygen in the air, we must ensure that all oxygen in the air reacts. This means that O_2 must be limited reactant while H_2 is excess reactant.

Hence at the end of reaction there is:

- Unreacted air (non-oxygen) components whose volume $= (25 - x)\text{cm}^3$
- Unreacted hydrogen gas whose volume $= (50 - 2x)\text{cm}^3$

But total volume left after the reaction$= 60\text{cm}^3$

Thus $25 - x + 50 - 2x = 60\,; 3x = 75 - 60 = 15\,; x = \frac{15}{3} = 5$

Whence volume of oxygen in 25cm^3 of the air $= 5\text{cm}^3$

Then $\%O_2\left(\frac{v}{v}\right) = \frac{\text{Volume of oxygen}}{\text{Volume of air}} \times 100\%$

$$= \frac{5}{25} \times 100\% = 20\%$$

Hence the percentage by volume of oxygen is 20%.

Question 40

CO_2 is obtained from decomposition of $CaCO_3$ according to the following equation:

$$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$$

From which mole ratio of $CaCO_3$ to CO_2 is $1:1$

That is $\frac{n_{CaCO_3}}{n_{CO_2}} = \frac{1}{1} = 1$ or $n_{CaCO_3} = n_{CO_2}$

But from $n = \frac{m}{M_r}$; Where $M_r(CaCO_3) = 100g/mol$

$$n_{CaCO_3} = \frac{75g}{100g/mol} = 0.75mol$$

Using $n = \frac{V}{GMV}$; $V = nGMV$ where at STP $GMV = 22.4dm^3$

Volume of CO_2 produced $= 0.75 \times 22.4dm^3 = 16.8\ dm^3$

Hence the volume of carbon dioxide obtained is $16.8\ dm^3$

Question 41

Let the volume of N_2O be $x\ cm^3$

And the volume of NO be $y\ cm^3$

Then $x + y = 60$…………………………….. (i)

N_2O reacts with H_2 according to the following equation:

$$N_2O + H_2 \rightarrow N_2 + H_2O$$

From which mole ratio of N_2O to N_2 is $1:1$

That is $n_{N_2O} = n_{N_2}$

But from Avogadro's law: Mole ratio = Volume ratio

Thus $V_{N_2O} = V_{N_2} = x\ cm^3$

Thus $x\ cm^3$ of N_2 is produced from the reaction between N_2 and H_2.

Also NO reacts with N_2 according to the following equation:

$$2NO + 2H_2 \rightarrow N_2 + 2H_2O$$

From which mole ratio of NO to N_2 is 2:1

That is $\frac{n_{NO}}{n_{N_2}} = \frac{v_{NO}}{v_{N_2}} = \frac{y}{v_{N_2}} = \frac{2}{1}$; $v_{N_2} = \frac{y}{2}cm^3$

Then total volume of N_2 produced $= x + \frac{y}{2} = 38$ …… (ii) (i)- (ii) gives: $\frac{y}{2} = 22\ ; y = 2 \times 22 = 44$

Substituting the value of y to (i) gives;

$$x + 44 = 60\ ;\ \ x = 60 - 44 = 16$$

Hence the composition is $16\ \mathrm{cm^3}$ of $\mathrm{N_2O}$ and $44\ \mathrm{cm^3}$ of NO

Question 42

$$\text{Density of solution} = \frac{\text{Mass of solution}}{\text{Volume of solution}}$$

From which;

$$\text{Mass of solution} = \text{Density of solution } \times \text{Volume of solution}$$

Where: Density of solution $= 1.08\ \mathrm{g/cm^3}$

Volume of solution $= 1\mathrm{L} = 1000\ \mathrm{cm^3}$

Thus mass of solution $= 1.08\mathrm{g/cm^3} \times 1000\mathrm{cm^3} = 1080\mathrm{g}$

$$\%\text{oxalic acid (solute)} = \frac{\text{Mass of oxalic acid}}{\text{Mass of solution}} \times 100\%$$

$$\%\text{oxalix acid (solute)} = \frac{3.24\mathrm{g}}{1080\mathrm{g}} \times 100\% = 0.3\% \left(\frac{\mathrm{m}}{\mathrm{m}}\right)$$

Hence the percentage of oxalic acid is $0.3\% \left(\frac{\mathrm{m}}{\mathrm{m}}\right)$

Question 43

(a)

$$\text{Mass concentration} = \text{Density of solution } \times \text{Mass percentage}$$

Where density of solution

$$= 1.8\mathrm{g/cm^3} = \frac{1.8\mathrm{g}}{0.001\mathrm{dm^3}} = 1800\ \mathrm{gdm^{-3}}$$

And mass percentage $= 24.5\%$

Thus mass concentration $= \frac{1800 \times 24.5}{100}\ \text{gdm}^{-3} = 441\ \text{gdm}^{-3}$

Using Molarity $= \frac{\text{Mass concentration in gdm}^{-3}}{\text{Molar mass}}$

Where molar mass of H_2SO_4

$$= \left((2 \times 1) + 32 + (4 \times 16)\right)\text{g/mol} = 98\text{g/mol}$$

Molarity $= \dfrac{441}{98}\text{M} = 4.5\text{M}$

Hence the molarity of the solution is 4.5M

(b)

Using $n = MV$;

Number of moles of $Ca(NO_3)_2$ in $100\ \text{cm}^3$ of 0.5M $Ca(NO_3)_2$ solution$= 0.5 \times \frac{100}{1000}\text{mol} = 0.05\text{mol}$

Number of moles of $Ca(NO_3)_2$ in $200\ \text{cm}^3$ of 1.25M $Ca(NO_3)_2$ solution$= 1.25 \times \frac{200}{1000}\text{mol} = 0.25\text{mol}$

Total number of moles of $Ca(NO_3)_2 = (0.05 + 0.25) = 0.3\text{mol}$

Total volume of the final solution

$$= (100 + 200)\text{cm}^3 = 300\text{cm}^3 = 0.3\text{dm}^3$$

Molarity of the final solution

$$= \frac{\text{Total number of moles}}{\text{Total volume of the solution}} = \frac{0.3\text{mol}}{0.3\text{dm}^3} = 1\text{moldm}^{-3}$$

Hence the molar concentration of the final solution is 1M

APPENDIX

LIST OF ATOMIC MASSES OF SELECTED ELEMENTS

Atomic Mass (Approximated)	Name chemical element	Symbol	Atomic number
1	Hydrogen	H	1
7	Lithium	Li	3
9	Beryllium	Be	4
11	Boron	B	5
12	Carbon	C	6
14	Nitrogen	N	7
16	Oxygen	O	8
19	Fluorine	F	9
23	Sodium	Na	11
24	Magnesium	Mg	12
27	Aluminum	Al	13
28	Silicon	Si	14
31	Phosphorus	P	15
32	Sulphur	S	16
35.5	Chlorine	Cl	17
39	Potassium	K	19
40	Calcium	Ca	20
52	Chromium	Cr	24
55	Manganese	Mn	25
56	Iron	Fe	26
64	Copper	Cu	29
65	Zinc	Zn	30
80	Bromine	Br	35
108	Silver	Ag	47
127	Iodine	I	53
137	Barium	Ba	56
207	Lead	Pb	82