THE MOLE

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Organizer

Objectives

By the end of this topic, the learner should be able to:

(a) Define the mole and relate it to relative atomic mass.
(b) Convert mass into moles, and moles to mass.
(c) Use data to determine the empirical and molecular formulae of compounds.
(d) Define the terms concentration, molarity and dilution of a solution.
(e) Define and prepare molar solutions.
(f) Carry out titrations and solve problems involving molar solutions.
(g) Write full formulae and ionic equations.
(h) Define molar gas volume and atomicity of gases.
(i) State Avogadros' and Gay Lussac's laws and carry out related calculations.

(j) Carry out calculations on molar gas volumes both at standard temperature and pressure and room temperature and pressure.

THE MOLE

The **mole** is the **SI unit for amount of substance.**

In chemistry, substances are called chemicals.

Relative Mass and the Avogadro's constant.

When the mass of any atom is compared to that of another, it is referred to as **relative atomic mass.**

The **Carbon-12** isotope is used as a reference for measuring relative atomic mass because it is a stable solid and a very common element.

The relative atomic mass of an element is defined as

$$\text{R.A.M.} = \frac{\text{Average mass of one atom of the element}}{\text{One twelfth } (\frac{1}{12}) \text{ of the mass of one atom of carbon-12}}$$

Since relative atomic mass is a ratio, it has no units.

Element	R.A.M	Element	R.A.M
Hydrogen	1	Magnesium	24
Carbon	12	Sulphur	32
Oxygen	16	Calcium	40
Sodium	23		

Relative Atomic masses of some selected elements

The **number of atoms** in **one relative atomic mass unit in grams of any element** has been established to be **6.023×10^{23}**. This number is referred to as **Avogadro's Constant, 'L'.**

A mole is the **amount of any substance** that contains Avogadro's number of particles **(6.023×10^{23} particles)** .

The unit "Mole" is used to **measure the amount of particles (atoms, molecules, ions, electrons etc.) of any substance.**

The mass in grams of one mole of a substance is referred to as **molar mass**

Inter Conversion of Molecules and Moles

Molecules contain more than one atoms of the same or different elements.

There are two atoms of oxygen in one molecule of oxygen. Therefore, the relative molecular mass of oxygen molecule is, $16 \times 2 = 32$. This implies that a mole of oxygen molecules has a mass of 32 g. This mass contains 6.023×10^{23} molecules of oxygen.

Since each molecule of oxygen has two atoms, then 1 mole of oxygen molecules would contain $2 \times 6.023 \times 10^{23} = 1.2046 \times 10^{24}$ atoms of oxygen. Similarly, 0.5 mole of oxygen molecules would contain: $6.023 \times 10^{23} \times 0.5 = 3.0115 \times 10^{23}$ molecules of oxygen and $3.0115 \times 10^{23} \times 2 = 6.023 \times 10^{23}$ atoms of oxygen.

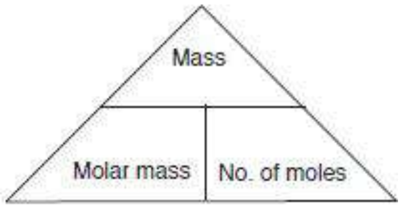
Inter conversion of Mass and Moles

(i) Relationship between mass of an element and the mole.

Given that one mole of an element has a mass equal to its relative atomic mass in grams, then it can be deduced that:

Mass of an element (in grams) = Molar mass × number of moles.

This formula can easily be remembered by using the triangle below:



$$\text{No of moles} = \frac{\text{Mass}}{\text{Molar mass}}$$

Key: • Horizontal line in triangle represents division sign.

- Vertical line in triangle represents multiplication sign.

Therefore, it would be seen that:

$$\text{Molar mass} = \frac{\text{Mass}}{\text{No of moles}}$$

$$\text{Mass} = \text{No. of moles} \times \text{molar mass.}$$

Worked Examples

1. **Calculate the mass of 0.4 moles of calcium (Ca = 40).**

Mass = molar mass × number of moles

= 0.4 × 40

= 16 g

2. **Determine the number of moles in 13.5 g of aluminium (A1 = 27)**

27 g = 1 mole

$$13.5\ g = \frac{1}{27} \times 13.5$$

$$= 0.5 \text{ moles.}$$

(ii) Relationship Between Relative Molecular Mass and the Mole

The **Relative molecular mass (for molecules**) of an element or compound is the **sum of all the relative atomic masses of the atoms in a molecule of the element or compound**.

For example, the relative molecular mass of chlorine Cl_2 is 35.5 × 2 = 71 g/mol

That of water, H_2O is (1×2) + 16 = 18g/mol.

Relative formula mass (for ionic substances) is the **sum of the relative atomic masses of all the atoms in a formula unit of a compound**.

For example, the relative formula mass of potassium sulphate K_2SO_4 is (39 × 2) + 32 + (16 × 4) = 174 g/mol.

NB: **One mole** of any substance has a **mass equal to the relative formula mass of that substance**. The substance could be made of atoms, ions or molecules.

2. Empirical and Molecular Formula

The **empirical formula** of a compound ***shows the simplest whole number ratio in which atoms combine to form the compound*.** The formula is determined from experimental data. It provides useful information about the chemical formula of a compound.

Empirical formula

Experiment to determine the formula of the compound formed when magnesium burns in air.

Weigh a clean dry crucible together with its lid. Clean about 15 cm length of magnesium ribbon thoroughly. Using a pencil, wind the magnesium ribbon into a coil. Place it inside the crucible and replace the lid. Weigh again the crucible together with the magnesium ribbon and lid. Heat the contents strongly for a few minutes, occasionally lifting the lid slightly using a pair of tongs..

When there are no more flare-ups, remove the lid and heat the crucible strongly. Remove the source of heat and allow the crucible to cool. When cold, replace the lid and weigh again.

Repeat the heating and cooling until a constant mass is obtained. Record your results in a table

Discussion Questions.

1. **Why is it necessary to clean the magnesium ribbon?**
   The magnesium ribbon is cleaned at the start of the experiment so as to remove any oxide film on it.
2. **Why was the lid kept on the crucible at first?**
   It is important to keep the lid in place to prevent any solid from escaping.
3. **What was the purpose of the occasional lifting of the lid?**
   It is necessary to lift the lid from time to time to allow in air.
4. **Explain why it is necessary to heat the crucible and its contents until there is no further change in mass.**
   The purpose of heating until a constant mass is obtained ensures that all the magnesium has reacted.
5. **What masses of magnesium and oxygen combine to form magnesium oxide?**

Sample Results

Mass of empty crucible + lid, (A)	19.52 g
Mass of crucible + lid + magnesium, (B)	20.36 g
Mass of crucible + lid + magnesium oxide, (C)	20.92 g
Mass of magnesium, (D) = (B – A)	0.84 g
Mass of magnesium oxide, (E) (C – A)	1.40 g

Mass of oxygen, (F = E – D)	0.56 g

6. **How many moles of atoms of each element reacted in this experiment?**
   Calculate the mole ratio of magnesium to oxygen in magnesium oxide.

The number of moles of atoms of magnesium, and oxygen which combine can be found by dividing the reacting masses of these elements by their respective relative atomic masses.

The formula of magnesium oxide can now be determined as follows:

Reacting elements:	**Magnesium**	**Oxygen**
Reacting masses (g):	0.84	0.56
Relative atomic mass:	24	16
No of moles =	$\frac{0.84}{24}$	$\frac{0.56}{16}$
Mole ratio	0.35	0.35
(divide by the smallest)	1	:1

7. **What is the simplest formula of magnesium oxide?**

It follows that, one mole of magnesium atoms combines with one mole of oxygen atoms. Therefore, The simplest formula of magnesium oxide is **MgO.**

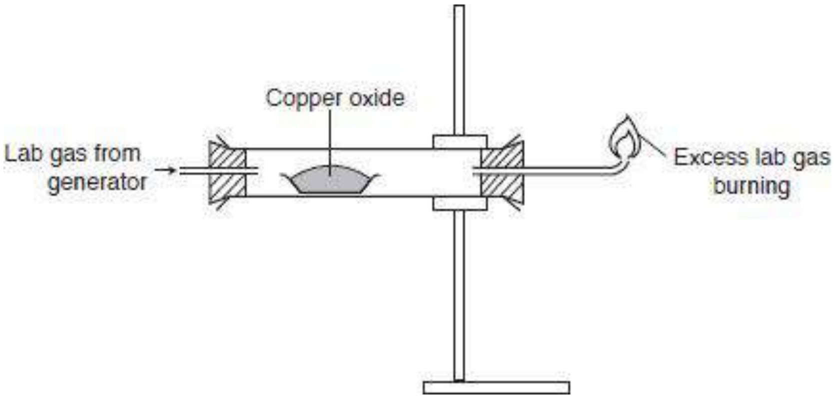
In this case, it also happens to be the chemical formula..

8. **State the possible sources of error in this experiment.**

Any variation from this ratio is due to experimental errors and side reaction such as the formation of magnesium nitride.

Experiment to determine the formula of the compound formed when copper combines with oxygen.

Weigh an empty porcelain boat. Place a small amount of copper(II) oxide into the porcelain boat and weigh again. Place the porcelain boat and its contents in a combustion tube as shown below. Pass a stream of laboratory gas through the tube for a short time. Light the gas at the end.



Start to heat the copper(II) oxide and record your observations. Continue heating the tube until there is no further change. Remove the source of heat but keep the lab gas flowing. When the tube has cooled, turn off the gas supply. Carefully remove the porcelain boat with the residue from the combustion tube and weigh. Repeat the procedure until a constant mass is obtained. Record the results in a table.

Discussion Questions

1. **Why is it necessary to ensure that all the air in the combustion tube is driven out before heating?**

Before heating, lab gas was passed through the tube to remove any traces of air in the tube

1. **How would you know that the reaction is complete?**

When all the black copper(II) oxide has been reduced to reddish brown metallic copper.

2. **What is the colour of the residue?**

In this experiment, black copper(II) oxide is reduced to reddish brown metallic copper using hydrogen gas. The oxygen combines with lab gas to form water and carbon(IV) oxide.

3. **Explain why it is necessary to keep the stream of lab gas on as the product is cooling.**

Metallic copper was cooled in a stream of lab gas to prevent re-oxidation of the hot metal by air.

5. **What other reducing agents can be used to remove oxygen from copper(II) oxide? Explain why some of these agents could not be used in this experiment.**

Ammonia gas can be used as the reducing agent in place of lab gas. Other reducing agents such as carbon and metals that are more reactive than copper can also be used to remove oxygen from copper(II) oxide but they are normally not used because it would be difficult to isolate the copper.

4. **Determine the formula of the copper oxide from these results.**

Mass of empty porcelain boat, (A)	15.6 g
Mass of porcelain boat + copper(II) oxide, (B)	19.1 g
Mass of porcelain boat + residue, (C)	18.4 g
Mass of copper(II) oxide, (D) = (B – A)	3.5 g
Mass of copper, (E) = (C – A)	2.8 g
Mass of oxygen, (F) = (D – E)	0.7 g

The formula can be determined as follows:

Element	**Copper**	**Oxygen**
Mass:	2.8 g	0.7 g
Relative atomic mass:	63.5	16
No. of moles:	$\frac{2.8}{63.5} = 0.044$	$\frac{0.7}{16} = 0.044$
Mole ratio	1	:1

One mole of copper combines with one mole of oxygen. The simplest formula is CuO. This is also the chemical formula of copper(II) oxide.

7. Determine the empirical formula using the percentage by mass of the combining elements.

Chemical formula is also derived from percentage composition of the constituent elements. Using the data for copper(II) oxide experiment, Percentage composition by mass of copper is: $\frac{\text{Mass of copper}}{\text{Mass of copper(II) oxide}} \times 100$

$$= \frac{2.8}{3.5} \times 100 = 80\%$$

Percentage composition by mass of oxygen is: $\frac{\text{Mass of oxygen}}{\text{Mass of copper oxide}} \times 100$

$$= \frac{0.7}{3.5} \times 100 = 20\%$$

The percentages are then taken to represent the actual masses of the elements in the compound.

The empirical formula is then determined as follows:

Element	**Copper**	**Oxygen**
Percentage composition:	80	20
Relative atomic mass:	63.5	16
No. of moles:	$\frac{80}{63.5}$	$\frac{20}{16}$
	= 1.25	= 1.25
Mole ratio:	1	:1

The empirical formula is CuO

Worked Example 1

An oxide of silicon was found to contain 47% by mass silicon. What is the empirical formula of the oxide? (Si = 28, O = 16).

Solution

Element	**Silicon (Si)**	**Oxygen (O)**
Percentage composition:	47	100 – 47 = 53
Relative atomic mass:	8	16
No. of moles $=\frac{\text{Composition \%}}{\text{Relative atomic mass}}$	$\frac{47}{28} = 1.68$	$\frac{53}{16} = 3.31$
Mole ratio:	$\frac{1.68}{1.68}$	$:\frac{3.31}{1.68}$
	= 1	:= 1.94
Simple whole No. ratio	1	:2

Therefore, the empirical formula of silicon oxide is SiO_2.

Worked Example 2

The percentage composition by mass of an oxide of iron is 70% iron and 30% oxygen. Determine its empirical formula. (Fe = 56, O = 16).

Solution

Element	**Iron(Fe)**	**Oxygen(O)**
Composition %:	70	**30**
Relative atomic mass:	56	**16**
No. of moles:	$\frac{70}{56} = 1.25$	$\frac{30}{16} = 1.875$

	$\frac{1.25}{1.25}$	$\frac{1.875}{1.25}$
Mole ratio:	1	:1.5

To make it a whole number, ratio multiply by 2

2 : 3

The empirical formula of the iron oxide is Fe_2O_3.

(b) Molecular Formula

Molecular formula shows the **actual number of each kind of atoms present in a molecule of the compound.**

If the empirical formula is known, then the molecular formula can be determined by the relationship.

(Mass of empirical formula)$_n$ = Molecular mass,

Where n is a whole number.
The molecular mass is **always a multiple of the empirical formula mass.**

Worked Example 1

A hydrocarbon was found to contain 92.3% carbon, and the remaining is hydrogen. If its molecular mass is 78, determine its molecular formula. (C = 12, H = 1).

Solution

Element	Carbon(C)	Hydrogen(H)
Percentage composition:	92.3	7.7
R.A.M:	12	1
No. of moles:	$\frac{92.3}{12}$	$\frac{7.7}{1}$
	= 7.7	:= 7.7
Mole ratio:	1	:1

The empirical formula is CH.

The molecular formula is determined as follows:

$$\begin{aligned} (CH)_n &= 78 \\ (12+1)n &= 78 \\ 13n &= 78 \\ n &= 6 \end{aligned}$$

Therefore, the molecular formula is C_6H_6

Worked Example 2

A compound of carbon, hydrogen and oxygen contains 54.55% carbon, 9.0% hydrogen, and 36.6% oxygen. If its relative molecular mass is 88, what is its molecular formula? (C = 12, O = 16, H = 1).

Solution

First obtain the empirical formula as follows:

Element	Carbon (C)		Hydrogen (H)		Oxygen (O)
Composition percentage:	54.55		9.09		36.36
R.A.M:	12		1		16
No. of moles:	$\frac{54.55}{12} = 4.55$		$\frac{9.09}{1} = 9.09$		$\frac{36.36}{16} = 2.27$
Mole ratio:	$\frac{4.55}{2.27}$	:	$\frac{9.09}{2.27}$	:	$\frac{2.27}{2.27}$
	2.004	:	4.004	:	1
Simple whole No. ratio:	2	:	4	:	1

The simplest formula is C_2H_4O.

Since the molecular mass is 88, then

$$(C_2H_4O)_n = 88$$

$$(12 \times 2) + (4 \times 1) + (16 \times 1)_n = 88$$

$$44n = 88$$

$$n = 2$$

The molecular formula is $C_4H_8O_2$.

Worked Example 3

When a certain hydrocarbon is burnt completely in excess oxygen, 5.28 g of carbon (IV) oxide and 2.16 g of water were formed. If the molecular mass of the hydrocarbon is 84, determine the molecular formula of the hydrocarbon.

Solution

This is an example where the mole ratio of the products can be used to determine the formula of the reactant.

Products:	**CO2**	**H20**
Mass:	5.28	2.16
Formula mass:	44	18
No. of moles:	$\frac{5.28}{44} = 0.12$	$\frac{2.16}{18} = 0.12$
Mole ratio:	1	:1

Therefore, it implies that only one mole of the carbon(VI) oxide and one mole of water were produced.

The empirical formula is obtained by working out the masses of carbon and hydrogen in carbon(IV) oxide and water respectively.

Products:	**Carbon (CO_2)**	**Water (H_2O)**
Mass:	$\frac{12}{44} \times 5.28 = 1.44$	$\frac{2}{18} \times 2.16 = 0.24$
No. of moles:	$\frac{1.44}{12} = 0.12$	$\frac{0.24}{1} = 0.24$
Mole ratio:	$\frac{0.12}{0.12}$	:$\frac{0.24}{0.12}$
	1	:1

Therefore, the empirical formula is CH_2.

Since the molecular mass is 84, then

$(CH_2)_n = 84$

$(12 + 2)_n = 84$

$14n = 84$

$n = 6$

The molecular formula of the hydrocarbon, is C_6H_{12}.

3. Molar Solutions

A molar solution is a solution that **contains one mole of a solute in one litre of the solution**.

Concentration of a Solution

Concentration of a solution is the **amount of a solute contained in a given volume of the solution.**

The concentration of a solution may be expressed in terms of **mass of solute in grams per given volume or number of moles of the solute per given volume**.

For **uniformity**, concentration is normally expressed in either **grams per litre** of solution **(g/dm³)** or in **moles per litre of solution (mol/dm³).**

Note: 1 litre = 1000 cm^3 = 1 dm^3

Example

Suppose 4.0 g of sodium hydroxide were dissolved in 400 cm^3 of distilled water then made up to 500 cm^3 of solution, determine the concentration in:

(a) grams.

(b) moles.

Solution

(a) The concentration is 4.0 g/500 cm^3 of solution.

Mass in 500 cm^3 = 4.0 g

$$\text{Mass in } 1000 \text{ cm}^3 = \frac{4.0 \text{ g}}{500 \text{ cm}^3} \times 1000 \text{ cm}^3$$
$$= 8.0 \text{ g}$$

In grams per litre, the concentration of the solution would be expressed as 8.0 g/dm^3

(b) The concentration is 4.0g/40g/mol. = 0.1 mole/500 cm^3 of solution.

Moles in 500 cm^3 = 0.1 mole

$$\text{Moles in } 1000 \text{ cm}^3 = \frac{0.1 \text{ mole}}{500 \text{ cm}^3} \times 1000 \text{ cm}^3$$
$$= 0.2 \text{ mole.}$$

Thus, the concentration in moles per dm-3 is 0.2 mole/dm^3.

It can be shown that 1 cm^3 of 4 g/500 cm^3 solution and 1 cm^3 of 8.0 g/1000 cm^3 contain the same mass of the solute.

(i) Mass in 500 cm^3 = 4.0 g

Mass in 1 cm^3 = $\frac{4.0 \text{ g}}{500 \text{ cm}^3}$ = 0.008 g cm^{-3}

(ii) Mass in 1000 cm^3 = 8.0 g

(Mass in 1 cm^3 = $\frac{8.0\ g}{1000\ cm^3}$ cm^3 = 0.008 gcm^{-3}

Thus, the concentration of a 500 cm^3 solution containing 4 g of solute is the same as the concentration of a 1000 cm^3 solution containing 8 g of solute.

Molarity of a Solution

The **molarity** of a solution is the **number of moles of the solute per litre of solution (moles/1000 cm^3)**. This is normally expressed as **mol dm^{-3}.**

Example

Suppose 71g of sodium sulphate are dissolved in enough water then made to one litre of solution. Determine the molarity of the solution formed. (Na = 23, S = 32, O = 16).

Mass of Na_2SO_4 dissolved = 71.0 g

Mass of 1 mole of Na_2SO_4 = 142 g

No. of moles in 1 dm^3 = $\frac{71\ g}{142\ g\ mol^{-1}}$ = 0.5 mole dm^{-3}

Molarity is 0.5 M Na_2SO_4.

Preparation of Molar Solutions

The apparatus used to prepare molar solutions include volumetric flasks, and measuring cylinders.

Preparing a solution involves dissolving the required mass of the solute in a little water say, 100 cm^3 of distilled water in a beaker. When all the solute has dissolved, the solution is transferred to the volumetric flask. The beaker is rinsed and the solution is transferred into the flask.

The solution is then made up to the mark on the narrow neck of the flask. The flask is then stoppered. The neck of the flask is narrow to ensure high accuracy of the instrument. The flask is then shaken to obtain a uniform solution.

Experiment: To prepare a molar solution of sodium hydroxide.

Weigh accurately 40.0 g of sodium hydroxide pellets, put them in 200 cm^3 of distilled water in a 250 cm^3 beaker and stir to dissolve. When all the solid has dissolved, transfer the solution into a one-litre volumetric flask. Rinse the beaker with some distilled water, and transfer the solution into the volumetric flask. Using distilled water in a wash bottle, make up the solution to the mark. Repeat the procedure using exactly 20.0 g and 10.0 g, of sodium hydroxide, prepare the solutions in 500 cm^3 and 250 cm^3 volumetric flasks respectively.

Questions

1. **Calculate the number of moles of sodium hydroxide in the:**
   (a) 1 litre solution.

Mass dissolved = 40.0 g
Molar mass of NaOH = 40.0 g mol^{-1}

No. of moles in 1 litre solution $= \frac{40.0\ g}{40.0\ g\ mol^{-1}}$

$= 1.0$ mole

(b) 500 cm^3 solution.

Mass dissolved = 20.0 g
Mass of NaOH = 40.0 g

No. of moles in 500 cm^3 $= \frac{20.0\ g}{40.0\ g\ mol^{-1}}$

$= 0.5$ mole.

(c) 250 cm^3 solution.(Na = 23, O = 16, H = 1).

Mass of NaOH dissolved = 10.0 g
Molar mass of NaOH = 40.0 g

No. of moles in 250 cm^3 $= \frac{10.0\ g}{40.0\ g\ mol^{-1}}$

$= 0.25$ mole

2. Determine the molarity of the solutions of sodium hydroxide in (1) (a), (b) and (c). Comment on the results.

(a) No. of moles in 1000 cm^3 = 1.0 mole dm^{-3}

Molarity is 1.0 M NaOH.

(b) No. of moles in 500 cm^3 = 0.5 mole

No. of moles in 1000 cm^3

$$= \frac{0.5\ g}{500\ cm^3} \times 1000\ cm^3 = 1.0\ mole\ dm^{-3}$$

Molarity is 1.0 M NaOH.

(c) No. of moles in 250 cm^3 = 0.25 mole

No. of moles in 1000 cm^3 = $\frac{0.25\ mole \times 1000\ cm^3}{250\ cm^3}$ = 1.0 mole dm^{-3}

Molarity is 1.0 M NaOH.

Dilution of a Solution

Dilution is a **process by which the concentration of a solution is lowered by adding more solvent into the solution.** During dilution, the **amount of solute remains the same as the volume of the solution increases.**

Experiment to dilute 2 M HCl

Measure accurately 25 cm^3 of 2 M HCl. Transfer the solution into a 250 cm^3 volumetric flask. Add distilled water to the acid and make up to 250 cm^3 of solution. Repeat the process using a 500 cm^3 volumetric flask instead of the 250 cm^3 flask.

Questions

1. **Determine the number of moles in the 25 cm^3 of the 2.0 M HCl.**

$$\text{Moles of HC1 in 1000 cm}^3 = 2.0 \text{ moles}$$
$$\text{Moles of HC1 in 25 cm}^3 = \frac{2.0 \text{ Moles}}{1000 \text{ cm}^3} \times 25 \text{ cm}^3$$
$$= 0.05 \text{ moles}$$

2. **Determine the molarity of the solution contained in the:**
   (i) 250 cm^3 flask.
   (ii) 500 cm^3 flask.

(a) Moles of HC1 in 250 cm^3 = 0.05 moles

$$\text{Moles in 1000 cm}^3 = \frac{0.05 \text{ Moles} \times 1000 \text{ cm}^3}{250 \text{ cm}^3}$$
$$= 0.2 \text{ mole}$$

Molarity of the solution is 0.2M HC1.

(b) Moles of HC1 in 500 cm^3 = 0.05 moles

$$\text{Moles in 1000 cm}^3 = \frac{0.05 \text{ Moles} \times 1000 \text{ cm}^2}{500 \text{ cm}^2}$$
$$= 0.1 \text{ mole}$$

Molarity of the solution is 0.1M HC1.

(iii) Comment on the results.

Although equal volumes of solution containing equal moles were diluted, the molarities of the resulting solutions are different. The solution to which less water was added is more concentrated than the one to which more water was added.

Solutions of different molarities may be prepared by adding water to equal volumes of solutions of the same concentration and making up to different volumes.

3. **Derive a general formula of preparing a less concentrated solution from a more concentrated solution.**

The molarity of the dilute solution (M_2) is obtained by multiplying the molarity of the concentrated solution (M_1) with the volume of the concentrated solution (V_l), and dividing by the final volume of the dilute solution (V_2);

Thus: $M_2 = \frac{M_1V_1}{V_2}$

Therefore, $M_2V_2 = M_1V_1$

The following worked examples show how this relationship is applied in calculations.

Example 1

Calculate the volume of a 5.0 M H_2SO_4 solution that will be required to make a 1000 cm^3 solution of 0.05 M H_2SO_4.

Solution

$M_1V_1 = M_2V_2$

M_1 = 5.0 M, V_1 = ? M_2 = 0.05 M, V_2 = 1000 cm^3

$$V_1 = \frac{M_2V_2}{M_1}$$

Substituting

$$= \frac{0.05 \times 1000 \text{ cm}^2}{5.0}$$
$$= 10 \text{ cm}^3$$

Example 2

When 50 cm^3 of 2M potassium hydroxide solution was diluted, the final concentration was O.1M. Calculate the volume of the diluted solution.

Solution:

$$M_1V_1 = M_2V_2$$

$M_1 = 2,\ V_1 = 50 \text{ cm}^3$

$M_2 = 0.1,\ V_2 = ?$

Substituting

$$V_2 = \frac{2 \times 50}{0.1}$$
$$V_2 = 1000 \text{ cm}^3$$

Example 3

Calculate the volume of 18 M sulphuric(IV) acid, H_2SO_4, that will be required to prepare 3.6 litres of 0.2 M sulphuric(VI) acid.

Solution

From $M_1V_1 = M_2V_2$

$$V_1 = \frac{M_2V_2}{M_1}$$

But, $M_2 = 0.2$ M

$V_2 = 3.6$ litres

$M_1 = 18$ M

$$\Rightarrow V = \frac{0.2 \times 3.6}{18}$$
$$= (0.2 \times 0.2) \text{ litres}$$
$$= 0.04 \text{ litres (Convert into cm}^3\text{)}$$
$$= 40 \text{ cm}^2$$

Example 4

A label on a bottle containing sulphuric(IV) acid has the following information:

- **Density = 1.836 g/cm^3**
- **Percentage purity = 98%**
- **Relative formula Mass = 98**

Calculate:

(a) **The concentration of the acid.**

(b) **The volume of the concentrated sulphuric(VI) acid that should be diluted to produce 2 litres of 2 M sulphuric(VI) acid.**

Solution:

(a) Mass of H_2SO_4 in $1cm^3 = \frac{1.836 \text{ g}}{1 \text{ cm}^3} \times 1cm^3$

$= 1.836$ g

Mass in 100 $cm^3 = 1.836 \times 1000$

$= 1836$ g

Mass of pure acid $= 1836 \text{ g} \times \frac{98}{100}$

$= 1799.28$ g

Molarity of acid $= \frac{\text{Mass per litre}}{\text{Relative formula Mass}} = \frac{1799.28}{98}$

$= 18.36$ M

(b) From the equations,

$M_1V_1 = M_2V_2$ $\quad V1 = \frac{M_2V_2}{M_1}$ $\quad$ Where, $M_2 = 2$, $V_2 = 2$, and $M_1 = 18.36$

Substituting: $\quad V_1 = \frac{2 \times 2}{18.36}$

$V_1 = 0.217864923$ litres $= 218$ cm^3

4. Stoichiometry of Chemical Equations

Stoichiometry is a **quantitative relationship between reactants and products in a chemical reaction.**

When particles combine in a chemical reaction, they do so in specific proportions. The proportion in which they combine is referred to as **mole ratio.**

In a balanced chemical equation, the quantities of reactants and products are in a whole number ratio. For example in the equation;

$$2Mg(s) + O_2(g) \longrightarrow 2MgO(s),$$

The reactants and product are in the ratio 2:1:2.

A chemical equation in which the reactants and products are in whole number ratios is called a **stoichiometric equation.**

Experiment to determine the equation for the reaction between iron metal and copper(II) sulphate solution.

Weigh accurately 0.56 g of freshly acquired iron filings and put them in a clean dry weighed beaker. Take about 20 cm^3 of 2 M copper(II) sulphate solution and transfer it into a clean boiling tube, and heat it until it is nearly boiling. Add the hot copper (II) sulphate solution to the iron filings in the beaker and stir the mixture to cool and record your observations.

Carefully decant as much of the liquid as possible, ensuring that no solid is lost. Wash the solid twice with distilled water. Place the beaker with the solid residue on a sand bath. When the solid is sufficiently dry remove the beaker from the sand bath and allow to cool. Weigh the beaker again with its contents. Record your results in a table.

Mass of iron filings	0.56 g
Mass of copper displaced	0.64 g

Discussion Questions

1. **What observations were made when iron filings and copper(II) sulphate reacted?**

When iron reacts with copper ions in solution, the products formed are brown copper metal and iron(II) sulphate in solution.

Copper(II) sulphate + Iron metal $\longrightarrow$ Iron(II) sulphate + Copper metal.

2. **Why was it necessary to use excess copper(II) sulphate?**

To ensure that all the iron metal used is reacted, blue copper(II) sulphate is used in excess, hence the solution remains blue when the reaction is completed.

3. **Calculate the:**

   (i) Number of moles of iron metal reacted.

Moles of iron used = $\frac{0.56\text{ g}}{56\text{ g mol}^{-1}}$ = 0.01 mole

   (ii) Number of moles of copper metal displaced from the solution. (Fe = 56, Cu = 63.5).

Moles of copper metal displaced = $\frac{1.64\text{ g}}{63.5\text{ g mole}}$ = 0.01007 mole

1. **Write the equation of the reaction between iron and copper ions in solution.**

From the masses of iron metal used, and the copper metal displaced from the solution, the reacting mole ratio of iron metal and copper ions in solution can be calculated.

Thus, Fe: Cu^{2+}

0.01: 0.01007

1.0: 1.007

Whole number mole ratio 1:1

Hence, one mole of iron atoms reacts with one mole of copper ions to produce one mole of atoms of copper metal. Since one mole of copper sulphate contains one mole of copper atoms, the equation of the reaction is

$$Fe(s) + CuSO_4(aq) \longrightarrow Cu(s) + FeSO_4(aq)$$

Experiment to determine the equation for the reaction between lead(II) nitrate and potassium iodide.

Take six test-tubes and label them, 1 to 6. Run 5 cm^3, of 1.0 M potassium iodide solution from a burette into each one of them. Add 1.0 cm^3 of 1.0 M lead(II) nitrate solution to the test-tube labelled 1, and stir the mixture well with a glass rod. Add about 5 drops of ethanol to the mixture, stir, and place it in a test-tube rack.

Add 1.5 cm^3, 2.0 cm^3, 2.5 cm^3, 3.0 cm^3 and 3.5 cm^3 of the 1.0 M lead(II) nitrate to the test-tubes labelled 2, 3, 4, 5, and 6 respectively. Add about 5 drops of ethanol to each test-tube, stir and allow to settle. Measure the height of the precipitate in each tube in (mm) and record the measurements in a table. Plot a graph of the heights of the precipitate against the volume of lead(II) nitrate solution added.

Sample results

Test tube number	1	2	3	4	5	6
Volume of 1 M $Pb(NO_3)_2$ (cm^3)	1.0	1.5	2.0	2.5	3.0	3.5
Height of precipitate (mm)	8	12	16	20	20	20

Discussion Questions

1. **What was observed on mixing the two solutions?**

When lead(II) nitrate solution reacted with potassium iodide solution, a bright-yellow precipitate of lead(II) iodide was formed. The insoluble lead(II) iodide settled at the bottom of the test-tube when allowed to rest.

2. **What was the purpose of adding ethanol to the mixture?**

Ethanol was added to the mixture to speed up the settling of the precipitate. Warming the mixture also hastens the settling process.

3. **Calculate:**

(a) Number of moles of KI in 5 cm³ of 1.0 M KI solution.

Moles of KI in 5.0 cm3 = $\frac{1.0 \times 5\ cm^3}{1000\ cm^3}$ = 0.005 moles

(b) Number of moles of $Pb(NO_3)_2$ which reacted completely with 5.0 cm^3 of 1.0 M KI.

Moles of $Pb(NO_3)_2$ that reacted completely, $\frac{1.0\ mole \times 2.5\ cm^3}{1000\ cm^3}$ = 0.0025 moles

4. **The heights of the precipitate remained constant in the test-tubes labelled 4, 5 and 6. Explain.**

The heights of the precipitate in test-tubes labeled 4 to 6, remained constant because lead(II) ions were in excess and all the iodide ions had reacted.

2. **How many moles of KI would react with one mole of lead(II) nitrate?**

Thus, lead ions and iodide ions reacted in the ratio 0.0025 : 0.005

i.e., $Pb^{2+} : I^-$

0.0025 : 0.005

Whole number ratio is 1 : 2

6. **Write:**

(a) Balanced chemical equation.

$$Pb(NO_3)_2(aq) + 2KI(aq) \longrightarrow PbI(s) + 2KNO_3(aq)$$

(b) Ionic equation for the reaction between lead(II) nitrate and potassium iodide.

Both the lead(II) ions (Pb^{2+}) and iodide ions (1^-) have undergone a chemical change. However nitrate ions and the potassium ions remained unchanged after the reaction. The ionic equation for the reaction can therefore be represented as:

$$Pb^{2+}(aq) + 2I^-(aq) \longrightarrow PbI_2(s)$$

Experiment: To determine the equation for the reaction between barium ions and carbonate ions.

Measure exactly 25.0 cm^3 of 0.2 M barium chloride solution, and place it in a beaker. Add 25 cm^3 of 0.2 M sodium carbonate solution. Stir the mixture and allow it to settle. Filter the solution, and dry the solid residue between a filter paper. Do not dry by heating. Transfer the dry solid onto a filter paper which you have weighed and recorded its weight. Record your results.

Discussion Questions

1. **What ions are present in:**

   (a) barium chloride solution?

   Ba2+ and Cl– ions

   (b) sodium carbonate solution?

   Na+ and CO_3^{2-} ions

3. **What observations were made when the two solutions were mixed? Explain.**

A White precipitate is formed.

On mixing the solutions, barium ions combine with carbonate ions to form insoluble barium carbonate. Sodium and chloride ions remain in solution unchanged.

4. **The solid residue is not dried by heating. Explain.**

The barium carbonate obtained should not be dried by heating because the compound easily decomposes on heating.

$$BaCO_3(s) \xrightarrow{\text{heat}} BaO(s) + CO_2(g)$$

5. **Calculate the number of moles of barium ions and carbonate ions used in the reaction. What assumptions have you made?**

Assuming that 25.0 cm^3 of 0.2M sodium carbonate reacted completely with 25.0 cm^3 of 0.2 M barium chloride., then

(a) Moles of barium chloride used

$$= 25\ cm^3 \times \frac{25\ cm^3 \times 0.2\ moles}{1000\ cm^3} = 0.005\ mole$$

(b) Moles of sodium carbonate used $= \frac{25\ cm^3 \times 0.2\ moles}{1000\ cm^3}$

$= 0.005$ mole.

6. **Calculate the number of moles of barium carbonate formed.**

Suppose 0.985 g of barium carbonate were formed, then:

Moles of $BaCO_3(s)$ formed $= \frac{0.985\ g}{197\ mol^{-1}}$

$= 0.005$ mole.

7. **What is the whole number ratio of barium carbonate formed to that of barium ions and carbonate ions used?**

1:1. One mole of barium chloride reacts with one mole of sodium carbonate to produce one mole of barium carbonate.

7. **Write the:**

   (i) stoichiometric equation.

The equation for the reaction is:

$$BaCl_2(aq) + Na_2CO_3(aq) \longrightarrow BaCO_3(s) + 2NaCl(aq)$$

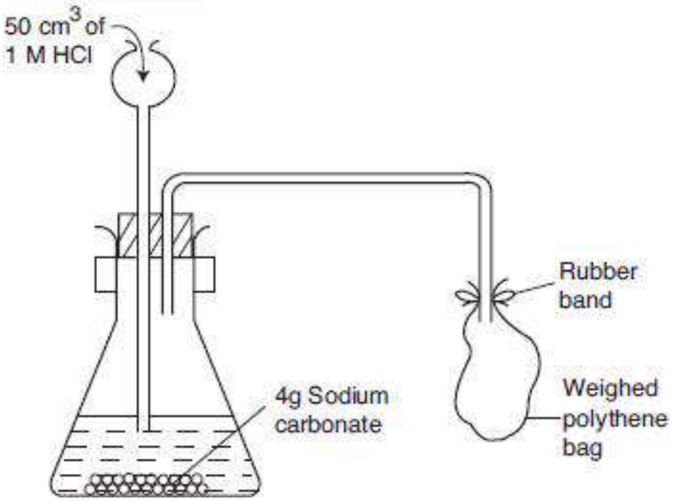
   (ii) ionic equation for the reaction.

The ionic equation for the reaction is:

$$Ba^{2+}(aq) + CO_3^{2-}(aq) \longrightarrow BaCO_3(s).$$

Experiment to determine the equation of reaction between hydrochloric acid and sodium carbonate.

Weigh a polythene bag and a rubber band and record their masses. Place 2.65 g of sodium carbonate in a conical flask. Stopper the flask with a cork fitted with a thistle funnel and delivery tube. Connect the delivery tube to the weighed polythene bag and tie the mouth of the bag tightly with the rubber band as shown below.



Measure exactly 50.0 cm^3 of 1M hydrochloric acid solution and transfer the acid to the thistle funnel. Run the acid from the thistle funnel into the conical flask. Collect the gas evolved in the bag. Remove the bag when the reaction has stopped. Weigh the bag with the gas again and record the mass.

Discussion Questions.

1. **Calculate:**
   (a) Moles of the acid used and Moles of the gas collected.

2. Write the equation for the reaction between hydrochloric acid and sodium carbonate.

$$\text{Moles of } H^+(aq) \text{ from HC1} = \frac{50\ cm^3 \times 1.0\ \text{mole}}{1000\ cm^3} = 0.05\ \text{mole}.$$

$$\text{Moles of } CO_3^{2-}(aq) \text{ from } Na_2CO_3 = \frac{2.65\ g}{106\ g\ mol^{-1}} = 0.025\ \text{mole}.$$

Assuming all the acid and carbonate reacted and the mass of carbon(IV) oxide collected was found to be 1.1g, then, moles of carbon(IV) oxide $= \frac{1.1\ g}{44\ mol^{-1}}$

= 0.025 mole.

Thus, 0.025 moles of CO_3^{2-}(aq) reacted with 0.05 moles of H^+(aq) to form 0.025 moles of carbon(IV) oxide.

The whole number mole ratio is: H^+	:	CO_3^{2-}	:	CO_2
0.05	:	0.025	:	0.025
2	:	1	:	2

The ionic equation is:

$$2H^+(aq) + CO_3^{2-}(aq) \longrightarrow CO_2(g) + H_2O(l)$$

and the stoichiometric equation is

$$2HCl(aq) + Na_2CO_3(aq) \longrightarrow 2NaCl(aq) + CO_2(g) + H2O(l)$$

5. Volumetric Analysis

Volumetric analysis, also called **titration,** is a method of quantitative chemical analysis in which the amount of a substance is determined by measuring volumes of solution. One solution of known concentration and volume is reacted carefully with another of unknown concentration to determine the reacting volumes. This data is used to determine the concentration of the said solution.

A solution whose exact concentration is known is called a **standard solution.**

The apparatus used in volumetric analysis are **pipettes** and **burettes.**

A pipette is designed to deliver a definite volume of a solution, e.g., 10.0 cm^3, 20.0 cm^3 or 25.0 cm^3.

Before use, the pipette **must be rinsed with the solution to be drawn.** It is then filled by sucking the liquid to **a few centimeters above the calibration mark.**

The solution is then allowed to **drain slowly under gravity until the meniscus is at the same level with the mark.** Sucking can be done using the mouth or a pipette filler. Pipette fillers **must** be used especially when solutions are toxic.

A burette is designed to deliver variable volumes of solution as needed during a titration. It is graduated from 0.0 to 50.0 cm^3 with unit intervals of 0.1cm^3.

Before use, a burette must be **rinsed with the solution to be used in it.** It is then filled carefully **beyond** the 0.0 mark and the level of the solution is **adjusted until the bottom of the meniscus** is at the same level with the graduation mark.

Titration

Titration is a quantitative analysis process using solutions. A solution of known concentration is added gradually to another solution of unknown concentration until the reaction between the two solutions is complete.

The point at which the reaction is complete is called the **end point**. An indicator is used to identify the **end point.**

The volume of the solution that run out of the burette in every titration experiment is known as a **titre.**

After the titration the data should be arranged in tabular form. The volume of the pipette used should be shown always.

The volume of pipette used .... cm^3.

Table of results of a titration experiment

Titration	1	2	3
Final burette reading (cm^3)			
Initial burette reading (cm^3)			
Volume (dilute) used (cm^3)			

Average volume used = ... cm^3.

Average titre = ... cm^3.

The titre should be recorded in **one or two decimal places. When in two decimal places, the last digit should be zero or five.**

The volumes used for the three experiments should be **consistent and have an accuracy of ± 0.2 cm^3.** The average titre is given by **adding the volumes of the consistent titres used divided by number of experiments done.**

Indicators are useful in pointing out the precise end-point of a titration. The choice of an indicator for every titration experiment therefore depends on its properties as shown below.

Indicator	Colour in Acid solution	Colour in Alkali solution	Colour in neutral solution
Phenolphthalein	Colourless	Pink	Colourless
Methyl orange	Pink	Yellow	Orange
Screened Methyl orange	Red	Green	Grey

Choice of Indicators

The table bellow is a guide on how to choose an appropriate indicator for different acid base titrations.

Titration	Suitable Indicator for use
Strong acid Strong alkali	Any indicator
Strong acid Weak base	Screened methyl orange or methyl orange
Weak acid Strong base	Phenolphthalein

The three basic Titrations are direct, back and redox Titrations.

Direct Titration

This is a type of titration that involves the addition of a standard solution to a fixed volume of the analyte (substance being anlaysed), or vice versa with the sole aim of determining either:

(a) Concentration.

(b) Relative atomic mass of the elements of the sample analyte or,

(c) The percentage composition of the constituent elements.

A good example is **a netralisation reaction**.

Experiment 1: What is the equation for the reaction between sodium hydroxide and hydrochloric acid?

Using a pipette, transfer 25 cm^3 of 0.1 M sodium hydroxide solution into a conical flask. Add 2 drops of phenolphthalein indicator. Fill a clean burette with 0.1 M hydrochloric acid and read the initial level of the acid accurately.

Run the acid solution in about 1 cm^3 portions, swirl the liquid in the flask after each addition of the acid until the colour of the solution changes to colourless permanently.

Record the final reading of the burette. Rinse the conical flask and repeat the experiment. Record the results in the table below.

The volume of the pipette used is .... cm^3.

Titration	1	2	3
Final burette reading (cm^3)			
Initial burette reading (cm^3)			
Volume acid used (cm^3)			

Average volume of acid used cm^3.

Discussion Questions

1. **What was the volume of hydrochloric acid that reacted with 25 cm^3 of 0.1 M sodium hydroxide solution?**

   Take the average titre as 25.0 cm3
2. **How many moles of hydrochloric acid are there in the average titre?**

Taking the average titre as 25.0 cm^3, therefore the number of moles of hydrochloric acid used are:

$$\frac{25}{1000} \times 0.1 = 0.0025$$

3. **How many moles of sodium hydroxide are there in 25 cm^3 of 0.1 M sodium hydroxide solution?**

Number of moles of sodium hydroxide used are:

$$\frac{25}{1000} \times 0.1 = 0.0025$$

4. **Write down the equation for the reaction between hydrochloric acid and sodium hydroxide.**

The mole ratio of acid to base is: 1:1, therefore, the equation is written as:

$$HCl(aq) + NaOH(aq) \longrightarrow NaCl(aq) + H_2O(1).$$

Mole ratio: 1:1.

- The reaction between an acid and an alkali forms a salt and water only. This kind of reaction is called **neutralisation.**
- The ionic equation is

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

Experiment 2: What is the equation for the reaction between sodium hydroxide and sulphuric(VI) acid?

Pipette 25.0 cm^3 of 0.1 M sodium hydroxide solution in to a conical flask. Add 2 drops of phenolphthalein indicator and titrate with 0.1 M sulphuric(VI) acid from a burette. Record the results as shownm

Volume of pipette used .... cm^3

Titration	1	2	3
Final burette reading (cm^3)			
Initial burette reading (cm^3)			
Volume (dilute) used (cm^3)			

Average volume acid used (cm^3)

Discussion Questions

1. **What volume of sulphuric(VI) acid neutralised 25 cm^3 of 0.1 M sodium hydroxide solution?**

   Taking the average titre as 12.5 cm3 of sulphuric(VI) acid

2. **How many moles of sulphuric(VI) acid are there in the average titre?**

   Moles of sulphuric(VI) acid = $\frac{125}{1000} \times 0.1 = 0.00125$

3. **How many moles of sodium hydroxide are there in 25 cm^3 of 0.1 M sodium hydroxide?**

   Moles of sodium hydroxide = $\frac{25}{1000} \times 0.1 = 0.0025$

4. **How many moles of sodium hydroxide react with 1 mole of sulphuric(VI) acid? Write the equation for the reaction.**

The mole ratio of sulphuric(VI) acid: sodium hydroxide is 0.00125: 0.0025 which is 1: 2. Therefore, equation for reaction is:

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

Ionic equation:

$$2H^+(aq) + 2OH^-(aq) \longrightarrow 2H_2O$$

5. **From experiment (1), and b2) compare the volumes of 0.1 M sulphuric(VI) acid and 0.1 M hydrochloric acid which neutralise 0.1 M sodium hydroxide solution in each case.**

Comparing experiment 1 and 2, the volume of 0.1 M hydrochloric acid is twice the volume of 0.1 M sulphuric acid that reacted with 0.1 M sodium hydroxide solution. This means that sulphuric acid is able to release twice the number of hydrogen ions as compared to hydrochloric acid.

- The number of hydrogen ions which can be produced by a molecule of an acid on ionising is referred to as the **basicity** of the acid. Thus, the basicity of HCl and HNO_3 is one (1) while that of H_2SO_4 and (HOOC–COOH) ethanedioic acid is two(2).

Name of acid	Formula	Number of replaceable hydrogen	Basicity
Hydrochloric acid	HC1	1	Monobasic
Nitric acid	HNO_3	1	Monobasic
Ethanoic acid	CH_3COOH	1	Monobasic
Sulphuric acid	H_2SO_4	2	Dibasic
Ethanedioic acid	HOOC–COOH	2	Dibasic
Phosphoric acid	H_3PO_4	3	Tribasic

For example:

$$HCl(aq) \xrightarrow{water} H^+(aq) + Cl^-(aq)$$

$$H_2SO_4(1) \xrightarrow{water} 2H^+(aq) + SO_4^{2-}(aq)$$

$$CH_3COOH(aq) \xrightarrow{water} H^+(aq) + CH_3COO^-(aq)$$

- In ethanoic acid only one hydrogen atom is replaceable, the one attached to O–H, the others are not replaceable.

Experiment 3: How can hydrochloric acid solution be standardised using sodium carbonate?

Put about 100 cm^3 of distilled water in 250 cm^3 volumetric flask. Measure 2 cm^3 of concentrated hydrochloric acid in a 10 ml measuring cylinder and transfer it to the flask. Add distilled water to make up to the 250 cm^3 mark. Rinse the burette with the acid. Pour the acid solution into the burette.

Weigh exactly 1.325 g of anhydrous sodium carbonate, dissolve in water and make up to 250 cm^3 in a flask. Transfer 25 cm^3 of the sodium carbonate solution using a pipette into a conical flask. Add 2 or 3 drops of methyl orange indicator and titrate with the hydrochloric acid solution until the indicator changes colour from yellow to pink. Record your data.

Volume of pipette used is **25 cm^3**

Titration	1	2	3
Final burette reading (cm^3)			
Initial burette reading (cm^3)			
Volume of acid used (cm^3)			

Discussion Questions

1. **What is the meaning of the term 'standardisation'?**

   Standardisation is the process by which an unknown concentration of a solution is determined by use of a standard solution in a titration experiment

2. **Determine the average volume of acid used**.

   Suppose the average volume of the acid used is 22.5 cm3

3. **Write the equation for the reaction between sodium carbonate and hydrochloric acid.:**

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

4. **Determine:**

   (a) The molarity of the sodium carbonate solution.

Relative formula mass of Na_2CO_3 is 106 g.

If 1.325 g of Na_2CO_3 was dissolved in 250 cm^3, then in 1000 cm^3, the mass would be:

$$\frac{1.325}{250} \times 1000 = 1.325 \times 4 = 5.3g$$

Hence, molarity $= \dfrac{\text{Mass in grams/litre}}{\text{Molar mass}} = \dfrac{5.3}{106} = 0.005\ M$

(b) The number of moles of sodium carbonate that reacted with the acid.

Moles of Na_2CO_3 in 25 cm^3 $= \frac{25}{1000} \times 0.005 = 0.00125$ mole.

4. **Determine the number of moles of acid used.**

Mole ratio Na_2CO_3: HCl is 1: 2.

Therefore, moles of HCl in 22.5 cm3 = 2 × 0.00125 = 0.0025 mole

6. **Calculate the concentration of the acid in:**

(a) mol dm.$^{-3}$

moles of HCl in 1000 cm^3 $= \frac{0.0025 \times 1000}{22.5} = 0.11$ mole.

Molarity of hydrochloric acid is 0.11 M.

(b) g dm.$^{-3}$

Molar mass of HCl is 36.5 g. Concentration of HCl in grams/litre is:

36.5 × 0.11 = 4.015 grams/litre = 4.0 g/L.

Back Titration

Back titration is a method of volumetric analysis used to **determine the concentration (or amount) of a reactant**.

It involves reacting quantities of the substance being analysed with an excess amount of a suitable reagent whose volume and concentration is known. The quantity of the excess is then measured through titration. The amount of the reagent that reacted with the analyte is determined by subtracting the excess amount from the initial. The data is then used to determine the concentration (or amount) of the analyte.

Experiment 1: What is the atomic mass of divalent metal M?

Weigh exactly 0.50 g of a divalent metal carbonate, MCO_3. Put the weighed carbonate in a conical flask. Add to it 30.00 cm^3 of 0.50 M hydrochloric acid. Add two to three drops of phenolphthalein in the resulting solution. Titrate the solution against a 1.0 M sodium hydroxide solution until the colour of the solution just turns pink permanently. Record your results in a table. Repeat the experiment three times to complete the table.

Test	1	2	3	4
Final burette reading (cm^3)	5.5	10.6	15.6	20.5
Initial burette reading (cm^3)	0.0	5.5	10.6	16.6
Volume of base used (cm^3)	5.5	5.1	5.0	4.9

Discussion Questions

1. **Calculate the average volume of base used to neutralise the excess acid.**

The average titre (base) used is determined by considering the values which are consistent to within ± 0.20 range only.

$$\text{Average volume of base used} = \frac{5.1 + 5.0 + 4.9}{3}\ \text{cm}^3$$
$$= 5.0\ \text{cm}^3$$

2. **Determine:**

(a) The number of moles of base used.

Volume of base that neutralised excess acid = 5.0 cm^3. Therefore, 1000 cm^3 of base contains 1.0 mole.

$$\therefore 5 \text{ cm}^3 \text{ of the base would contain} = \frac{1.0 \text{ mole}}{1000 \text{ cm}^3} \times 5 \text{ cm}^3$$
$$= 0.005 \text{ mole.}$$

(b) (i) The volume of acid that was neutralised by the base.

The equation for the neutralisation reaction is:

$$HCl\ (aq) + NaOH\ (aq) \longrightarrow NaCl\ (aq) + H_2O\ (1)$$

Thus, the reacting mole ratio of acid to base, HCl: NaOH is 1: 1. Hence, moles of excess acid neutralised was 0.005 mole. But, the molarity of the acid was 0.50 M HCl and 0.50 moles was contained in a 1000 cm^3.

$$0.005 \text{ mole would be contained in} = \frac{1000 \text{ cm}^3 \times 0.005 \text{ moles}}{0.50 \text{ mole}}$$
$$= 10.0 \text{ cm}^3$$

Thus, volume of excess hydrochloric acid after reaction with the carbonate is 10.0 cm^3.

This means only 20.0 cm^3 of the 30.0 cm^3 of acid added to the carbonate to completion.

(ii) The number of moles of acid that reacted with the carbonate.

1000 cm^3 of acid contained = 0.50 mole.

$$20 \text{ cm}^3 \text{ of acid would contain} = \frac{0.50 \text{ mole} \times 20 \text{ cm}^3}{1000 \text{ cm}^3}$$
$$= 0.01 \text{ mole.}$$

3. (i) Write the equation for the reaction between the acid and carbonate.

$$MCO_3\ (s) + 2HCl(aq) \longrightarrow MCl(aq) + CO_2\ (g) + H_2O\ (l)$$
$$1 \quad : \quad 2$$

(ii) Determine the number of moles of the carbonate in the sample.

The reacting mole ratio of the carbonate to the acid is 1 : 2

MCO3 : 2HCl

1 : 2

Thus moles of the carbonate that reacted with 0.10 mole of the acid is $\frac{1}{2}$ of 0.10 mole.

$$\text{Mole of carbonate} = \frac{\text{moles of acid}}{2}$$
$$= \frac{0.10 \text{ mole}}{2} = 0.005 \text{ mole.}$$

(iii) Determine the relative formula mass of the carbonate.

Mass of 0.005 mole = 0.50 g

$$\text{Mass of 1.0 moles} = \frac{0.50 \text{ g}}{0.005 \text{ mole}} = 100 \text{ g}$$

$$\text{R.F.M. } (MCO_3) = 100 \text{ g}$$

(iv) Determine the relative atomic mass of M.

$$R.F.M.\ (MCO_3) = 100\ g$$
$$M + 12 + 48 = 100\ g$$
$$R.A.M.\ of\ M = (100 - 60)\ g = 40\ g$$

Alternatively,

$$R.F.M.\ of\ carbonate \times 0.005\ mole = 0.50\ g$$
$$0.005 \times (M + 12 + 48) = 0.50\ g$$
$$0.005\ M + 0.30\ g = (0.50 - 0.30)$$
$$0.005\ M = 0.20\ g$$
$$M = \frac{0.20}{0.005} = 40$$
$$R.A.M.\ of\ M = 40g$$

Redox Titrations.

Titrations involving redox reactions do not require indicators because the colour of some of the reagents change when their oxidation states change.

Examples of redox titrations involve the use of potassium manganate(VII) and potassium dichromate(VI). Manganate(VII) ions are purple in colour. When reduced to manganese(II), the solution becomes colourless.

$$\underset{Purple}{Mn^{7+}(aq) + 5\ e^-} \longrightarrow \underset{Colourless}{Mn^{2+}\ (aq)}$$

Dichromate(VI) ions are orange in colour. When reduced to chromium(III) ions, the solution becomes green.

$$\underset{Orange}{Cr^{6+}\ (aq) + 3\ e^-} \longrightarrow \underset{Green}{Cr^{3+}\ (aq)}$$

Redox titrations are used to standardise solutions and to determine the purity of compounds. Potassium manganate(VII) and potassium dichromate(VI) solutions **should be acidified before use in titrations. This is done to ensure complete reduction to manganese(II) and chromate(III) ions, respectively.**

Experiment 1: How can potassium manganate(VII) solution be standardised using iron(II) salt?

Weigh accurately 9.8 g of the iron(II) salt, $FeSO_4.(NH_4)_2SO_4.6H_2O$) and dissolve in distilled water that has been boiled and cooled to remove air. Add more distilled water to make up to 250 cm^3.

Transfer 25.0 cm^3 of the solution into a conical flask using a pipette. Add 10 cm^3 of dilute sulphuric acid and titrate with potassium manganate(VII) from a burette until a persistent pink colour appears. Record the volume of potassium manganate(VII) used. Repeat the experiment twice and record your results in a table

Titration	1	2	3
Final burette readings (cm^3)			
Initial burette reading (cm^3)			
Volume of $KMnO_4$ used (cm^3)			

Discussion Questions

1. **Determine the average volume of potassium manganate solution.**
   Suppose the average titre is 24.0 cm3.

2. Calculate the molarity of the iron(II) salt solution.

Molar mass of the iron(II) salt = 392 g.

Number of moles of iron(II) ions in 250 cm^3 solution = $\frac{9.8 \text{ g}}{392 \text{ mol}^{-}} = 0.025$ mole

Moles in 1000 cm^3 = $\frac{0.025 \text{ mole} \times 1000 \text{ cm}^3}{250 \text{ cm}^3}$ = 0.1 mole

Thus molarity of the iron(II) salt is 0.1 M.

3. Calculate the number of moles of iron(II) ions used.

Number of moles of iron(II) ions in 25 cm^3.

Solution, = $\frac{25.0 \text{ g} \times 0.1}{1000}$ = 0.0025 mole.

4. Given the equation of the reaction is:

$$MnO_4^-(aq) + 8H^+(aq) + 5Fe^{2+}(aq) \longrightarrow Mn^{2+}(aq) + 5Fe^{3+}(aq) + 4H_2O(l)$$

Determine:

(a) The number of moles of MnO_4^- ions in 24.0 cm^3 solution.

Reacting mole ratio $MnO_4^- : Fe^{2+}$

1 : 5

Therefore, moles of MnO_4^- in 24.0 cm^3, = 0.0025 × $\frac{1}{5}$

= 0.0005 mole.

(b) **The molarity of the $KMnO_4$(aq) solution**.

$$\text{Molarity of } KMnO_4 = \frac{0.005 \text{ moles} \times 100 \text{ cm}^3}{24.0 \text{ cm}^3}$$

$$= 0.020833 \text{ M} = 0.02 \text{ M}$$

Experiment 2: What is the amount of water of crystallisation in ammonium iron (II) sulphate?

Ammonium iron(II) sulphate crystals have the following formula $(NH_4)_2SO_4$. $FeSO_4$. nH_2O. Weigh accurately 8.8 g of the salt and dissolve in 50 cm^3 of 2.0 M sulphuric(IV) acid. Make up to 250 cm^3 of solution with distilled water. Pipette 25.0 cm^3 of this solution and pour it into a conical flask. Titrate this against the acidified manganate(VII) used in experiment 1 until the solution becomes colourless. Record your results in a table.

Titration	1	2	3
Final burette readings (cm^3)			
Initial burette reading (cm^3)			
Volume of $KMnO_4$ used (cm^3)			

Discussion Questions

1. Calculate the:

(a) Average volume of $KMnO_4$ solution used.

Suppose the average volume of the 0.02 M KMnO4 solution used was 22.5 cm3

(b) Number of moles of $KMnO_4$ that reacted.

Moles in 1000 cm^3 = 0.02 mole.

$$\text{Number of moles in } 22.5\text{ cm}^3 = \frac{0.020\text{ mole}}{1000\text{ cm}^3} \times 22.5\text{ cm}^3$$
$$= 0.00045\text{ mole.}$$

2. **Using the following ionic equation determine the number of moles of iron(II) salt in 25 cm³ of the solution used.**

$$MnO_4^-(aq) + 8H^+(aq) \longrightarrow Mn^{2+}(aq) + 5Fe^{3+}(aq) + 4H_2O(l).$$

But, from the ionic equation, the reacting mole ratio of MnO_4^- : Fe^{2+} is 1 : 5.

thus, moles of iron(II) ions reacted = 0.00045 × 5 mole
= 0.00225 mole.

3. **Determine the molarity of the iron(II) salt solution.**

0.00225 mole were contained in 25 cm³.

That is, 25 cm³ contained = 0.00225 mole.

$$250\text{ cm}^3 \text{ contained} = \frac{0.00225 \times 250}{25}$$
$$= 0.0225\text{ mole.}$$

Hence, molarity of iron(II) salt is determined as follows:

250 cm³ contains 0.0225 mole

$$1000\text{ cm}^3 \text{ contains} = \frac{0.0225\text{ mole} \times 1000\text{ cm}^3}{250\text{ m}^3} = 0.09\text{ mole.}$$

4. **Determine:**

   (a) The value of 'n' in the formula, $(NH_4)_2SO_4.FeSO_4.nH_2O$.

250 cm³ solution contains 8.8 g

$$1000 \text{ solution contains } \frac{8.8\text{ g} \times 1000\text{ cm}^3}{250\text{ cm}^3} = 35.2\text{ g}$$

Mass of 0.09 mole = 35.2 g

$$\text{Mass of } 1.0\text{ mole} = \frac{35.2\text{ g} \times 1.0\text{ mole}}{0.09\text{ mole}}$$
$$= 391.1\text{g}$$

To find 'n', $36 + 32 + 64 + 56 + 32 + 64 + 18n = 391.1$

$$284 + 18n = 391.1$$
$$18n = 391.1 - 284$$
$$n = \frac{107.11}{18} = 5.9506 \simeq 6.0$$

   (b) The formula of the iron(II) salt.

The formula of the salt is: $(NH_4)_2SO_4.FeSO_4.6H_2O$

6. Atomicity and Molar Gas Volume

Atomicity

Atomicity is the number of atoms in one molecule of an element.

For example, oxygen (O_2), hydrogen (H_2) and nitrogen (N_2) have two atoms per molecule, thus they are **diatomic.**

Ozone (O_3) has three atoms in a molecule and is therefore **triatomic.**

The noble gases are said to have an atomicity of one (**monoatomic**) even though they are not molecular.

Atomicity of some selected gaseous elements.

Gas	Chemical formula	Atomicity
Helium	He	1
Argon	Ar	1
Neon	Ne	1
Nitrogen	N2	2
Hydrogen	H2	2
Oxygen	O2	2
Chlorine	Cl2	2
Ozone	O3	3
Bromine	Br2	2

Molar Gas Volume

The volume occupied by one mole of any gas at given temperature and pressure is called the **molar gas volume.** Its value at s.t.p is **22.4 dm³**. At **r.t.p**, the molar gas volume is **24.0 dm³**.

Combining Volumes of Gases

The relationship between reacting volumes of gases is summarised by Gay **Lussac's** Law: *When gases react, they do so in volumes that bear a simple ratio to one another and to the volumes of the product if gaseous, temperature and pressure remaining constant*.

For example:

$$\begin{array}{lllcl} \text{Nitrogen} & + & \text{Hydrogen} & \longrightarrow & \text{Ammonia Gas} \\ 10\ cm^3 & & 30\ cm^3 & & 20\ cm^3 \\ 1\ \text{vol} & & 3\ \text{vol} & & 2\ \text{vol} \end{array}$$

Example 2

Carbon(II) oxide gas + Oxygen gas ⟶ Carbon(IV) oxide gas

$2CO(g) + O_2(g)$	⟶	$CO_2(g)$
20 cm^3	10 cm^3	10 cm^3
2 vol	1 vol	1 vol

Example 3

Hydrogen gas + Oxygen gas ⟶ Steam

$2H_2(g)$	$+ O_2(g)$ ⟶	$2H_2O(g)$
30 cm^3	15 cm^3	30 cm^3
2 vol	1 vol	2 vol

Avogadros' Law also states that ***Equal volumes of gases will contain equal number of molecules.***

The number of molecules per mole of any gas is **6.02 × 10^{23}** and occupies **a volume of 22.4 dm^3 at s.t.p**. The number **6.02 × 10^{23} is** known **as the Avogadro's number.**

Avogadro's law implies that volume ratios of reacting gases can be used interchangeably with mole ratios when gases react. This means that reacting volumes of gases can be used to determine the equation of reacting gases provided the volumes are measured under the same conditions of temperature and pressure.

Worked Examples

1. **In an experiment, 20 cm^3 of sulphur(II) oxide are found to react completely with 10 cm^3 of oxygen to produce 30 cm^3 of sulphur(IV) oxide. Determine the equation for the reaction.**

Sulphur(IV) oxide + Oxygen ⟶ Sulphur(VI) oxide

gas	gas ⟶	gas
$SO_2(g)$	$+ O_2(g)$ ⟶	$SO_3(g)$
20 cm^3	10 cm^3	20 cm^3
2 vol	2 vol	2 vol
2 moles	1 mole	2 moles

So the balanced equation is:

$$2SO_2(g) + O_2(g) \longrightarrow 2SO_3(g)$$

2. **When 100 cm^3 of a gaseous hydrocarbon(C_xH_y) burn in 400 cm^3 of oxygen, 100 cm^3 of oxygen is unused, 200 cm^3 of carbon(IV) oxide and 200 cm^3 of steam are formed. Deduce the equation for the reaction and the formula of the hydrocarbon.**

Solution

$C_xH_y(g)$ +	$O_2(g)$ ⟶	$CO_2(g)$ +	$H_2O(g)$
100 cm^3	300 cm^3	200 cm^3	200 cm^3
1 vol	3 vol	2 vol	2 vol
1 mol	3 mol	2 mol	2 mol

The formula deduced from reacting volumes is worked out as follows:

$$C_xH_y(g) + O_2(g) \longrightarrow 2CO_2(g) + 2H_2O(g)$$

Writing a balanced equation from the above:

$$C_2H_4(g) + 3O_2(g) \longrightarrow 2CO_2(g) + 2H_2O(g)$$

The formula of the hydrocarbon is derived on the basis of the following facts:

(i) In a balanced chemical equation, the number of each kind of atoms on the reactant side must be equal to that in the product side.

(ii) Hydrocarbons are compounds consisting of carbon and hydrogen only.

Review Exercise 2

Revision Questions

1. 2006 Q 8 (P1)

When 94.5g of hydrated barium hydroxide, $Ba(OH)_2.nH_2O$ were heated to constant mass, 51.3g of anhydrous barium hydroxide were obtained. Determine the empirical formula of the hydrated barium hydroxide. (3 marks)

2. 2007 Q 20

An alkanol has the following composition by mass: hydrogen 13.5%, oxygen 21.6% and carbon 64.9%.
Determine the empirical formula of the alcohol (C=12.0; H=1.0) =16.0).

(2 marks)

3. 2007 Q 22

6.84g of aluminium sulphate were dissolve in 150cm^3 of water. Calculate the molar concentration of the sulphate ions in the solution.
(Relative formula mass of aluminium sulphate is 342) (3 marks)

4. 2008 Q 2

When a hydrated sample of calcium sulphate $CaSO_4.\mathbf{x}H_2O$ was heated until all the water was lost, the following data recorded;

Mass of crucible	=	30.296 g
Mass of crucible +hydrated salt	=	33.111 g
Mass of crucible + anhydrous salt	=	32.781 g

Determine the empirical formula of the hydrated salt.
(Relative formula mass of $CaSO_4$ =136, H_2O =18). (3 marks)

5. 2008 Q 5

Phosphoric acid is manufactured from calcium phosphate according to the following equation.

$$Ca_3(PO_4)_2(s) + 3H_2SO_4(l) \longrightarrow 2H_3PO_4(aq) + 3CaSO_4(s)$$

Calculate the mass (in kg) of phosphoric acid that would be obtained if 155 kg of calcium phosphate reacted completely with the acid
(Ca=40, P=31, S=32, O=16, H=1) (2 marks)

6. 2008 Q 27

In an experiment to determine the percentage of magnesium hydroxide in an anti-acid, a solution containing 0.50 g of the anti-acid was neutralized by 23.0 cm^3 of 0.010M hydrochloric acid. (Relative formula mass of magnesium hydroxide =58)
Determine the:
(a) Mass of magnesium hydroxide in the anti-acid; (2 marks)
(b) Percentage of magnesium hydroxide in the anti-acid. (1 mark)

7. 2009 Q 13 P1, 2016 Q7 P1
When 8.53g of sodium nitrate were heated in an open tube, the mass of oxygen gas produced was 0. 83g.Given the equation of the reaction is:

$$2NaNO_3(s) \xrightarrow{heat} 2NaNO_2(s) + O_2(g)$$

Calculate the percentage of sodium nitrate that was converted to sodium nitrite.
(Na= 23.0, N = 14.0, O = 16.0) (3 marks)

8. 2010 Q 6 P1
Aluminium oxide reacts with both acids and bases.
(a) Write an equation for the reaction between aluminium oxide and hydrochloric acid. (1 mark)

(b) Using the equation in (a) above, calculate the number of moles of hydrochloric acid that would react completely with 153.0g of aluminium oxide. (Al = 27.0, O= 16.0) (2 marks)

9. 2010 Q 8 P1
The pressure of nitrogen gas contained in a 1 dm^3 cylinder at -196 °C was 10^7 Pascals.
Calculate the:
(a) Volume of the gas at 25 °C and 10^5 Pascals. (1½ marks)
(b) Mass of nitrogen gas. (Molar volume of gas is 24 dm^3, N = 14.0) (1 ½ marks)

10. 2010 Q 17 P1
Analysis of a compound showed that it had the following composition:
69.42% carbon, 4.13% hydrogen and the rest oxygen.

(a) Determine the empirical formula of the compound. (C = 12.0, H = 1.0, O = 16.0) (2 marks)
(b) If the mass of one mole of the compound is 242, determine its molecular formula. (1 mark)

11. 2010 Q 3 P2, 2016 Q1 P2
Use the information in the table below to answer the questions that follow. The letters do not represent the actual symbols of the elements.

Element	Atomic number	Melting point (°C)
R	11	97.8
S	12	650.0
T	15	44.0
U	17	-102
V	18	-189
W	19	64.0

(a) Give the reasons why the melting point of:
  (i) **S** is higher than that of **R**; (1 mark)
  (ii) **V** is lower than that of **U**. (2 marks)

(b) How does the reactivity of **W** with chlorine compare with that of **R** with chlorine? Explain.
(2 marks)

(c) Write an equation for the reaction between T and excess oxygen. (1 mark)

(d) When 1.15g of **R** were reacted with water, 600cm^3 of gas was produced. Determine the relative atomic mass of R. (Molar gas volume = 24000 cm^3) (3 marks)

(e) Give **one** use of element **V**. (1 mark)

12. 2011 Q 7 P1
When lead (II) nitrate is heated, one of the products is a brown gas.

(a) Write the equation of the reaction that occurs. (1 mark)

(b) If 0.290 dm^3 of the brow n gas was produced, calculate the mass of the lead (II) nitrate that was heated. (R.F.M of lead (II) nitrate = 331; Molar gas volume = 24 dm^3).

(2 marks)

13. 2011 Q 25 P1
(a) State the Gay Lussac's Law. (1 mark)

(b) 10 cm^3 of a gaseous hydrocarbon, C_2H_x required 30cm^3 of oxygen for complete combustion. If steam and 20 cm^3 of carbon (IV) oxide were produced, what is the value of **x**?
(2 marks)

14. 2011 Q 26
The data given below was recorded when metal **M** was completely burnt in air. **M** is not the actual symbol of the metal. (R.A.M; M=56, O=16)

Mass of empty crucible and lid =10.240g
Mass of crucible, lid and metal M =10.352g
Mass of crucible, lid and metal oxide = 10.400g

(a) Determine the mass of:

(i) Metal M (½ mark)
(ii) Oxygen (½ mark)

(b) Determine the empirical formula of the metal oxide. (2 marks)

15. 2012 Q8 P1
$10cm^3$ of concentrated sulphuric (VI) acid was diluted to $100\ cm^3$. $10\ cm^3$ of the resulting solution was neutralised by $36\ cm^3$ of 0.1M sodium hydroxide solution. Determine the mass of sulphuric (VI) acid that was in the concentrated acid. (S = 32.0; H= 1.0; O = 16.0).
(3 marks)

16. 2012 Q11 P1
The empirical formula of **A** is CH_2Br. Given that 0.470g of **A** occupies a volume of $56\ cm^3$ at 546K and 1 atmosphere pressure, determine its molecular formula.
(H = 1.0, C= 12.0, Br = 80.0, molar gas volume at STP = $22.4\ dm^3$) (3 marks)

17. 2012 Q23 P1
Describe how the percentage by mass of copper in copper carbonate can be determined. (3 marks)

18. 2013 Q23 P1
When $15\ cm^3$ of a gaseous hydrocarbon, P, was burnt in $100cm^3$ of oxygen, the resulting gaseous mixture occupied $70\ cm^3$ at room temperature and pressure.
When the gaseous mixture was passed through potassium hydroxide solution, its volume decreased to $25\ cm^3$.

(a) What volume of oxygen was used during the reaction? (1 mark)
(b) Determine the molecular formula of the hydrocarbon (2 marks)

19. 2013 Q24 P1
A solution was made by dissolving 8.2g of calcium nitrate to give 2 litres of solution. (Ca= 40.0; N=14.0; O= 16.0)
Determine the concentration of nitrate ions in moles per litre. (3 marks)

20. 2013 Q1 P2
The grid given below represents part of the periodic table. Study it and answer the questions that follow. The letters do not represent the actual symbol of the element.

M				N	P	T		
R								

(a) Select a letter which represents an element that losses electrons most readily. Give a reason for your answer. (2 marks)

(b) Explain why the atomic radius of P is found to be smaller than that of N (2 marks)

(c) Element M reacts with water at room temperature to produce 0.2 dm^3 of gas. Determine the mass of M which was reacted with water.
(Molar gas volume at room temperature is 24 dm^3; Relative atomic mass of M=7) (3 marks)

21. 2014 Q13 P1
100cm^3 of 0.05 M sulphuric (VI) acid were placed in a flask and a small quantity of anhydrous sodium carbonate added. The mixture was boiled to expel all the carbon (IV) oxide. 25cm^3 of the resulting solution required 18 cm^3 of 0.1 M sodium hydroxide solution to neutralize it. Calculate the mass of sodium carbonate added.
(Na = 23.0; O=16.0; C=12.0) (3 marks)

22. 2014 Q25 P1
An organic compound had the following composition 37.21% carbon,
7.75% hydrogen and the rest chlorine. Determine the molecular formula of the compound, given that the molecular mass of the compound is 65.
(C=12.0; H=1.0; Cl=35.5) (3 marks)

23. 2015 Q5 P1
Calculate the mass of Zinc oxide that will just neutralize dilute nitric (V) acid containing 12.6 g of nitric (V) acid in water. (Zn = 65.0; O =16.0, H = 1.0, N = 14.0). (3 marks)

24. 2015 Q26 P1
A hydrocarbon contains 14.5% of hydrogen. If the molar mass of the hydrocarbon is 56, determine the molecular formula of the hydrocarbon.
(C = 12.0; H = 1.0) (3 marks)

25. 2015 Q1b P2
Zinc oxide can be obtained by heating zinc nitrate. A student heated 5.76 g of zinc nitrate.

(a) Write an equation for the reaction that occurred. (1 mark)

(b) Calculate the total volume of gases produced. (Molar gas volume is 24 dm^3; Zn = 65.4; O = 16.0; N = 14.0). (4 marks)

(c) Identify the element that is reduced when zinc nitrate is heated. Give a reason. (2 marks)

26. 2017 P1 Q4.
The empirical formula of lead(II) oxide was determined by passing excess dry hydrogen gas over 6.69g of heated lead(II) oxide.

(a) What was the purpose of using excess dry hydrogen gas? (2 marks)

(b) The mass of lead was found to be 6.21g. Determine the empirical formula of the oxide. (Pb = 207.0 0 = 16.0) (2 marks)

27. 2017 P1 Q10.

20 cm^3 of ethanoic acid was diluted to 400 cm^3 of solution. Calculate the concentration of the solution in moles per litre. (C = 12.0; H = 1.0; 0 =16.0) (Density of ethanoic acid = 1.05 g/cm^3) (3 marks)

28. 2017 P2 Q4.

W is a colourless aqueous solution with the following properties:

I. It turns blue litmus paper red.
II. On addition of cleaned magnesium ribbon, it gives off a gas that burns with a pop sound.
III. On addition of powered sodium carbonate, it gives off a gas which forms a precipitate with calcium hydroxide solution.
IV. When warmed with copper(II) oxide powder, a blue solution is obtained but no gas is given off.
V. On addition of aqueous barium chloride, a white precipitate is obtained.

(a) (i) State what properties (I) and (III) indicate about the nature of W. (1 mark)
(ii) Give the identity of W. (1 mark)
(iii) Name the colourless solution formed in (II) and (III). (2 marks)
(iv) Write an ionic equation for the reaction indicated in (V). (1 mark)

(b) Element V conducts electricity and melts at 933 K. When chlorine gas is passed over heated V, it forms a vapour that solidifies on cooling. The solid chloride dissolves in water to form an acidic solution. The chloride vapour has a relative molecular mass of 267 and contains 19.75% of V. At a higher temperature, it dissociates to a compound of relative molecular mass 133.5. When aqueous sodium hydroxide is added to the aqueous solution of the chloride, a white precipitate is formed which dissolves in excess alkali. (V = 27.0; CI = 35.5)

(i) Determine the:
I. empirical formula; (2 marks)
II. molecular formula. (2 marks)

(ii) Draw the structure of the chloride vapour and label the bonds. (1 mark)

(iii) Write an equation for the reaction that form a white precipitate with sodium hydroxide. (1 mark)

29. 2017 P2 Q5

(a) When 0.048 g of magnesium was reacted with excess dilute hydrochloric acid at room temperature and pressure, 50 cm^3 of hydrogen gas was collected. (Mg = 24.0; Molar gas volume = 24.0 dm^3)

(i) Draw a diagram of the apparatus used to carry out the experiment described above. (3 marks)

(ii) Write the equation for the reaction. (1 mark)

(iii) Calculate the volume of hydrogen gas produced. (2 marks)

(iv) Calculate the volume of 0.1 M hydrochloric acid required to react with 0.048 g of magnesium. (3 marks)

30. 2018 P1 Q 7.
30.0 cm^3 of aqueous sodium hydroxide containing 8.0 g per litre of sodium hydroxide were completely neutralised by 0.294 g of a dibasic acid. Determine the relative formula mass of the dibasic acid. (Na = 23.0; O = 16.0; H 1.0) (3 marks)

31. 2018 P1 Q28.
Distinguish between empirical and molecular formula of a compound. (1 mark)

32. 2019 P1 Q7.
A solution contains 40.3g of substance XOH per litre. 250.0 cm^3 of this solution required 30.0cm^3 of 0.3M sulphuric(VI)acid for complete neutralisation.

(a) Calculate the number of moles of XOH that reacted. (½ mark)
(b) Determine the relative atomic mass of X. (1½ mark)

33. 2019 P1 Q 13.
5 g of calcium carbonate was strongly heated to a constant mass.
Calculate the mass of the solid residue formed. (Ca = 40.0; C = 12.0; 0 = 16.0). (2 marks)

34. 2019 P1 Q 29.
Name an appropriate apparatus that is used to prepare standard solutions in the laboratory.
(1 mark)