MOLES AND EQUATIONS

Introduction:
In chemistry, the mole serves as a fundamental unit for quantifying the amount of substance. Grasping the concept of the mole is vital for understanding chemical equations, as it enables scientists to predict how substances will react in specific proportions.

Chemical equation illustrates a chemical reaction through symbols and formulas, ensuring that the number of atoms for each element remains consistent on both sides. By utilizing the mole concept, chemists can accurately determine the precise amounts of reactants required and the products generated,

Comprehending moles and chemical equations is crucial across various fields, including industrial chemistry, pharmaceuticals, and environmental science. This understanding facilitates accurate chemical measurements and promotes efficient resource utilization.

Sub topic 1.1 Masses of Atoms and Molecules, Accurate Relative Atomic Masses

Atoms and molecules serve as the essential building blocks of matter, and understanding their masses is vital for grasping chemical reactions and properties.

1. Atomic Mass (Mass of an Atom)

Atom is the smallest electrically neutral particle of an element that takes part in chemical reaction. The fundamental particles of an atom include protons, neutrons and electrons. Protons are positively charged; electrons are negatively charged and neutrons have no charge

Particle	Charge	Mass
Proton	$+$	1
Neutron	No charge	1
Electron	$-$	$\frac{1}{1836}$

A molecule is the smallest particle of an element or compound that can exist in a free and separate state.
Examples of molecules of elements include hydrogen (H_2), nitrogen (N_2), oxygen (O_2), chlorine (Cl_2), and fluorine (F_2)

The atomic mass of an element refers to the mass of a single atom, primarily determined by the total number of protons and neutrons in the nucleus. Electrons, having a negligible mass, contribute very little to the overall atomic mass.
Atomic mass is measured in atomic mass units (amu) or unified atomic mass units (u). Specifically, 1 amu is defined as $\frac{1}{12}th$ the mass of a carbon-12 atom, which is equivalent to 1.66054×10^{-27} kg.

Therefore, when we state that an oxygen atom has an atomic mass of 16 amu, it indicates that it is roughly 16 times heavier than $\frac{1}{12}$th of a carbon-12 atom.

Examples of Atomic Masses from the Periodic Table

- *Hydrogen (H) has an atomic mass of 1, consisting of 1 proton and no neutrons.*
- *Carbon (C) has an atomic mass of 12, with 6 protons and 6 neutrons.*
- *Oxygen (O) has an atomic mass of 16, made up of 8 protons and 8 neutrons.*
- *Sodium (Na) has an atomic mass of 23, which includes 11 protons and 12 neutrons.*

It's important to note that elements exist as mixtures of isotopes. As a result, the mass number listed on the periodic table typically represents a weighted average of all isotopes, leading us to the concept of relative atomic mass.

2. Isotopes

Isotopes are atoms of the same element that have the same number of protons but different numbers of neutrons.
This variation changes the atomic mass but does not significantly affect chemical properties, as chemical behavior depends primarily on the number of electrons, which remain the same in isotopes of an element.

Notation for Isotopes

Isotopes are written in the form:

For example, the three naturally occurring isotopes of hydrogen and carbon are:

- *Protium:* $^{1}_{1}H$
- *Deuterium:* $^{2}_{1}H$
- *Tritium:* $^{3}_{1}H$
- *Carbon 12:* $^{12}_{6}C$
- *Carbon 13:* $^{13}_{6}C$
- *Carbon 14:* $^{14}_{6}C$

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3. Relative Atomic Mass

There are at present 118 different elements known. The atoms of these elements differ in mass because of the different numbers of protons, neutrons and electrons they contain. The actual mass of one atom is very small.

Such small quantities are not easy to work with, a scale called the **relative atomic mass** scale is used. In this scale, an atom of carbon is given a relative atomic mass, of 12. All other atoms of the other elements are given a relative atomic mass compared to that of carbon.

Relative atomic mass, is the average mass of the isotopes of an element compared to $\frac{1}{12}$**th** of the mass of an atom of ^{12}C.

$$R.A.M = \frac{MASS\ OF\ AN\ ATOM\ OF\ AN\ ELEMENT}{\frac{1}{12}x MASS\ OF\ AN\ ATOM\ OF\ CARBON\ 12}$$

Example

1. The relative atomic mass of an element is determined by comparing the mass of one of its atoms to $\frac{1}{12}$th of the mass of a carbon-12 atom. Given that the mass of an oxygen atom is 2.66×10^{-23} g and the mass of a carbon-12 atom is 1.99×10^{-23}g, calculate:
The **relative atomic mass** of oxygen.

Solution

From **R.A.M** = $\frac{MASS\ OF\ AN\ ATOM\ OF\ AN\ ELEMENT}{\frac{1}{12}x MASS\ OF\ AN\ ATOM\ OF\ CARBON\ 12}$

$$R.A.M = \frac{2.66x10^{-23}}{\frac{1}{12}x1.99x10^{-23}}$$

$$R.A.M = 16$$

<u>Therefore, relative mass of oxygen is 16</u>

2. A metal recycling company processes scrap metal to extract pure iron for manufacturing. To verify the purity of an iron sample, a laboratory technician must calculate its relative atomic mass. Given that:

- The mass of **one iron atom** is **9.27×10^{-23}g**.
- The mass of **one carbon-12 atom** is **1.99×10^{-23}g**.

Tasks:

a) Calculate the ***relative atomic mass of iron****.*

Since elements can exist as mixtures of isotopes, their atomic mass is not a whole number. The ***relative atomic mass*** of an element is the ***weighted average mass*** of all naturally occurring isotopes, taking into account their ***abundances***.

$$\text{Relative atomic mass} = \frac{(m1 \times \%A1) + (m2 \times \%A2) + (m3 \times \%A3)}{100}$$

Where:
- m_1, m_2, m_3 = *mass of each isotope*
- A_1, A_2, A_3 = *percentage abundance of each isotope*

Examples

1. Chlorine has two naturally occurring isotopes:
   - *Chlorine-35 (^{35}Cl): Mass = 35 amu, Abundance = 75%*
   - *Chlorine-37 (^{37}Cl): Mass = 37 amu, Abundance = 25%*

$$\textit{Relative atomic mass of chlorine} = \frac{(35 \times 75) + (37 \times 25)}{100}$$

$$\frac{2625 + 925}{100}$$

$$= 35.5$$

*Thus, **the relative atomic mass of chlorine is 35.5 amu**, which is why the periodic table lists **Cl = 35.5**.*

2. Magnesium has three naturally occurring isotopes as
Magnesium-24 (^{24}Mg): Mass = 24 amu, Abundance = 79%
Magnesium-25 (^{25}Mg): Mass = 25 amu, Abundance = 10%
Magnesium-26 (^{26}Mg): Mass = 26 amu, Abundance = 11%

2. Calculate the relative atomic mass of magnesium

$$\textit{Relative atomic mass of magnesium} = \frac{(24 \times 79) + (25 \times 10) + (26 \times 11)}{100}$$

$$= \frac{1896 + 250 + 286}{100}$$

$$= 24.32 \text{ amu}$$

*Thus, **the relative atomic mass of magnesium is 24.32 amu**. which is why the periodic table lists **Mg = 24**.*

Determine Accurate Relative Atomic Mass Using Given Isotopic Data.

Determining the accurate relative atomic mass of an element is fundamental in understanding its chemical behavior and properties. This process involves calculating a weighted average of the masses of the element's naturally occurring

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isotopes, each contributing according to its fractional abundance. One of the key techniques employed in obtaining the precise isotopic data necessary for these calculations is mass spectrometry. A mass spectrometer ionizes a sample and separates the ions based on their mass-to-charge ratios, allowing scientists to measure both the masses and the relative abundances of the isotopes with exceptional accuracy.

Mass spectrometry

Is a technique used to identify the compounds present in the sample by vapourising the sample into mixture of gaseous ions (with or without fragmentation), sorting/separating the deflecting gaseous ions in the presence of magnetic field based on their mass to charge ratio (m/z) and recording the relative abundance of each ion.

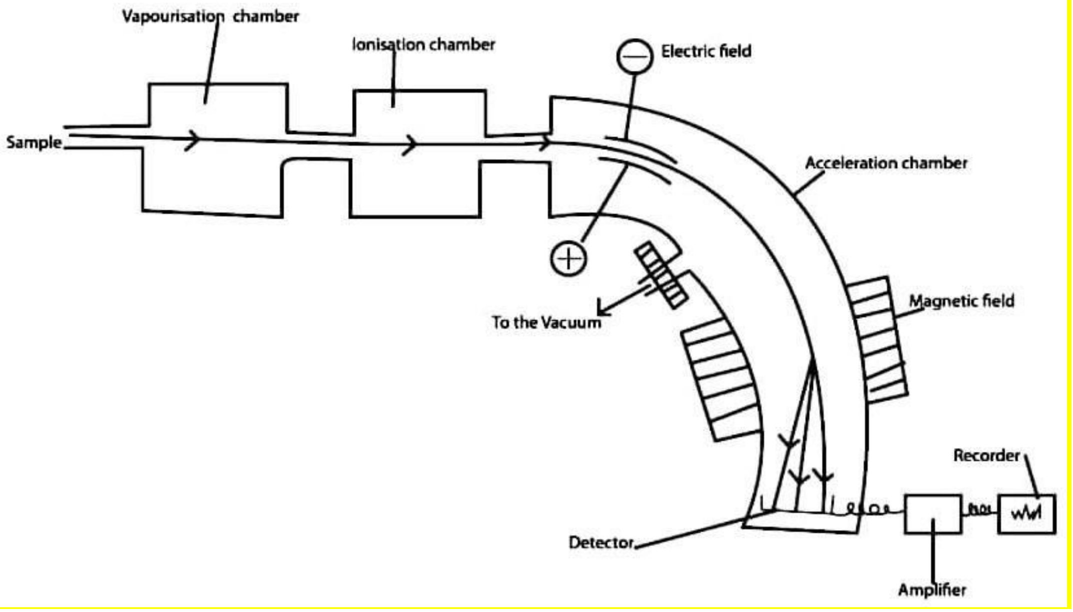
The earliest experiments of mass spectroscopy by J. J. Thomson used a stream of positive ions from a discharge tube, which were deflected by parallel electric and magnetic fields at right angles to the beam. Each type of ion formed a parabolic trace on a photographic plate (a mass spectrograph). In modern instruments, the ions are produced by ionizing the gas with electrons.

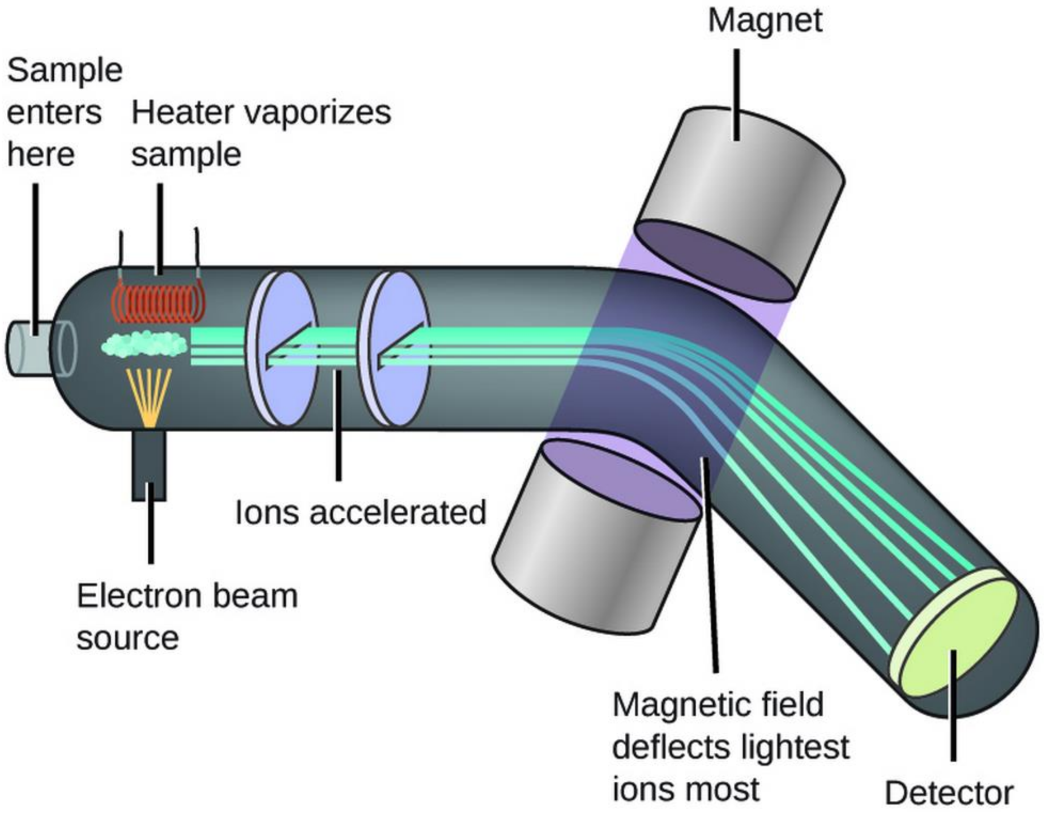
The essential parts and operation of a modern mass spectrometer

A mass spectrometer is usually used during mass spectroscopy. It is an instrument for producing ions in a gas and analyzing them according to their charge/mass ratio.

It mainly consists of the vapourisation chamber, ionisation chamber, acceleration chamber, deflection chamber, the detector, amplifier and recorder.

Diagram showing the essential parts of a modern mass spectrometer





Operation of a modern mass spectrometer

The whole apparatus is evacuated of air particles that may produce unnecessary ions.

1. Vapourisation chamber

The sample is heated by a tungsten filament to vapourise it. The vapourised sample is then introduced in the **ionisation chamber**.

2. The ionization chamber. In this chamber, the vapourised sample is subjected to a beam of fast-moving electrons emitted by a hot filament. These electrons bombard the atoms of the vapourised sample from which 1 or 2 electrons are removed to form positive ions.

$$X_{(g)} \longrightarrow X^{+}_{(g)} + e \quad \textit{for 1 electron}$$
$$X_{(g)} \longrightarrow X^{2+}_{(g)} + e \quad \textit{for 2 electrons}$$

3. The acceleration chamber.

The positive ions produced from the ionisation chamber are accelerated by a strong electric field of varying potentials such that only positive ions with the same velocity/kinetic energy but with different mass-charge ratio will pass through to the magnetic field.

4. The deflection chamber. The magnetic field then deflects the ions according to their mass-charge ratio. The strength of the magnetic field is varied such that ions of the same mass-charge ratio are focused onto the detector.

5. The detector, amplifier and recorder

The ions are changed into sizeable electric currents which are sent to the amplifier for fine tuning and then to the recorder where they are recorded into peaks that show relative intensities of ions of a particular mass falling at the detector at that instant. A mass spectrum is thus obtained consisting of a series of peaks of variable intensity to which mass/charge ratio (m/e) values can be assigned.

If the sample introduced was that of an element, its relative atomic mass (R.A.M) can be obtained by the formula;

$$R.A.M = \sum \text{Relative isotopic mass x proportion of isotope}$$

Follow the link: video on operation of mass spectrometer
https://youtu.be/Az2XfgBhP00?si=4BlMDlHRVYyRzzx5

APPLICATIONS OF MASS SPECTROMETER IN OUR DAILY LIFE

*The **mass spectrometer** is a powerful analytical tool used in various fields of daily life, from healthcare to environmental monitoring. Here are some key applications:*

1. Environmental Monitoring

In environmental analysis, mass spectrometry plays a pivotal role in detecting and quantifying pollutants. It's employed to analyses air, water, and soil samples for traces of organic and inorganic compounds. The sensitivity of mass spectrometry allows for the identification of contaminants at very low concentration levels, making it an essential tool for environmental protection and compliance with regulatory standards.

2. Medical and Healthcare Applications

Mass spectrometry has become an indispensable tool in medical research and diagnostics. It's used for a broad range of applications from identifying biomarkers for diseases to drug discovery and development. One of its significant contributions is in proteomics, where mass spectrometry analyses the structure and function of proteins involved in various diseases. Moreover, it's increasingly used in clinical laboratories for toxicology tests, therapeutic drug monitoring, and metabolic screening. The ability to rapidly and accurately measure drug levels in a patient's blood can be critical for personalised medicine, ensuring drugs are both effective and safe.

3. Personalized Nutrition and Agriculture

Ensuring the safety and quality of food is a major concern worldwide. Mass spectrometry assists in this area by detecting contaminants, such as pesticides, toxins, and allergens, in food products. It's also employed to authenticate food by verifying its origin and composition, which is paramount in preventing food fraud.

The mass spectrum

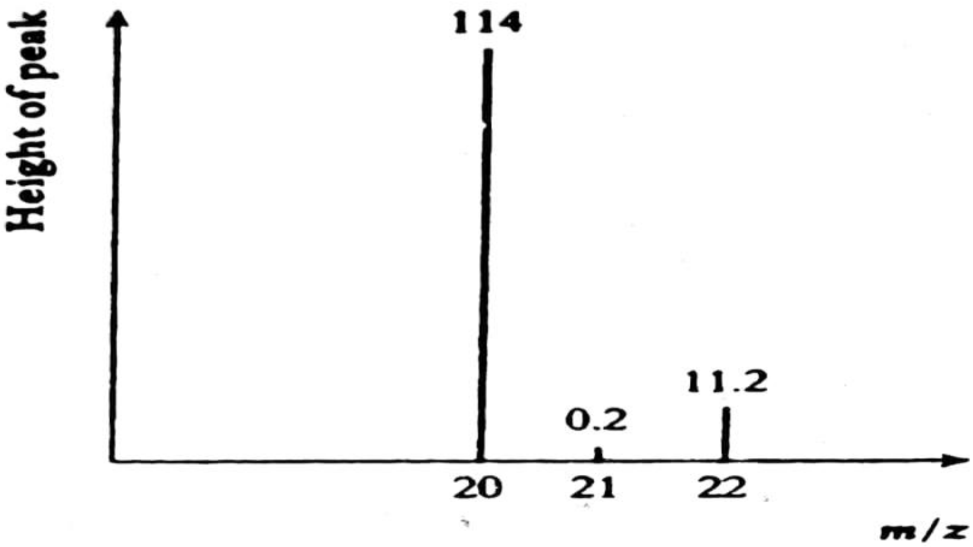
A mass spectrum is a plot of percentage abundance or relative intensity against mass to charge ratio of the ions separated in a mass spectrometer.
The mass to charge ratio is however numerically equal to the mass of the ion since most of the ions formed in a mass spectrometer have a single charge.

Interpreting a mass spectrum

The mass spectrum is usually has vertical lines and each vertical line represents an ion having a specific mass-to-charge ratio⁄and the length of the line indicates the relative abundance of the ion.

***Relative abundance** refers to the relative intensity of each isotope of an element represented as a ratio or percentage.*
The tallest line is assigned an abundance of 100 and is referred to as the **base peak.** The intensities of the other ions are measured relative to this line.

Consider the mass spectrum for neon below



According to this spectrum, there are three stable isotopes of neon. These include neon- 20(^{20}Ne), neon-21(^{21}Ne) and neon-22 (^{22}Ne). The peaks at 20, 21 and 22 are due to the ions $^{20}Ne^+$, $^{21}Ne^+$ and $^{22}Ne^+$ respectively. These ions are formed by bombardment of the gaseous atoms by electrons as shown below;

$$^{20}Ne_{(g)} + e \longrightarrow {}^{20}Ne^+_{(g)}$$
$$^{21}Ne_{(g)} + e \longrightarrow {}^{21}Ne^+_{(g)}$$
$$^{22}Ne_{(g)} + e \longrightarrow {}^{22}Ne^+_{(g)}$$

The height of each peak is related to the relative abundance of each isotope of neon. $^{20}Ne^+$has the highest peak so ^{20}Ne is the most abundant isotope. $^{21}Ne^+$has the lowest peak so ^{21}Ne is the least abundant isotope. The relative atomic mass of neon can be obtained using the mass spectrum. The percentage abundances (proportion of isotope) should be obtained first.

Total height = 114 + 0.2 + 11.2 = 125.4

Percentage abundance of Neon – 20 = $\frac{114}{125.4}x100$ = *90.91*

Percentage abundance of Neon – 21 = $\frac{0.2}{125.4}x100$ = *0.16*

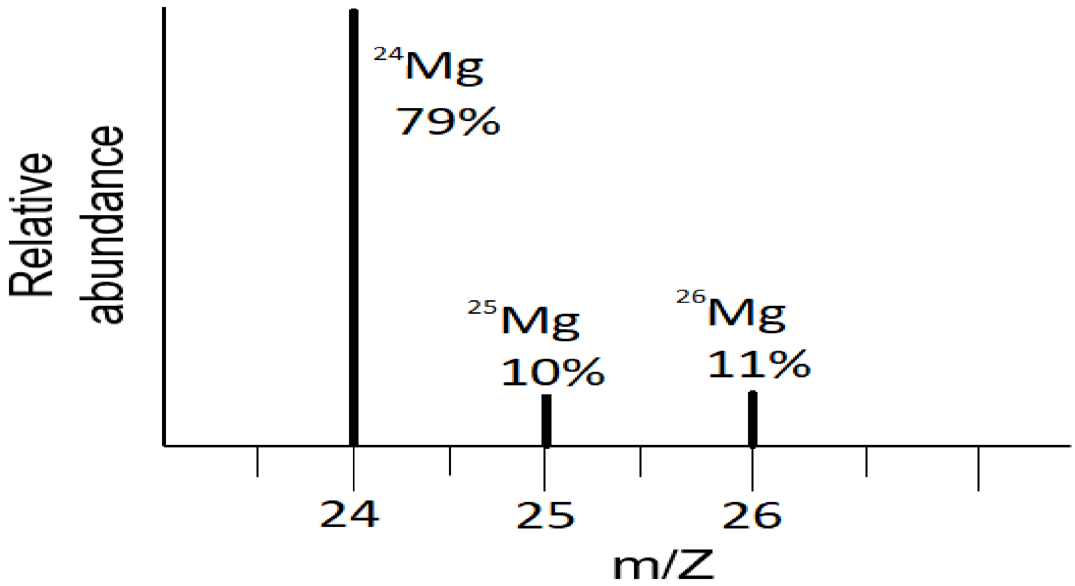
Percentage abundance of Neon – 22 = $\frac{11.2}{125.4}x100$ = *8.93*

R.A.M = Σ Relative isotopic mass x proportion of isotope

R.A.M = $(20x\frac{90.91}{100}+21x\frac{0.16}{100}+22x\frac{8.93}{100})$

R.A.M = 20.18

Consider mass spectrum for magnesium below



The spectrum has three peaks at mass to charge ratios, 24, 25, and 26. The peaks correspond to the ions of magnesium, $^{24}Mg^+$, $^{25}Mg^+$ *and* $^{26}Mg^+$ *with* $^{24}Mg^+$ *having the highest peak.*

To determine the relative atomic mass of magnesium, the peaks are used as below

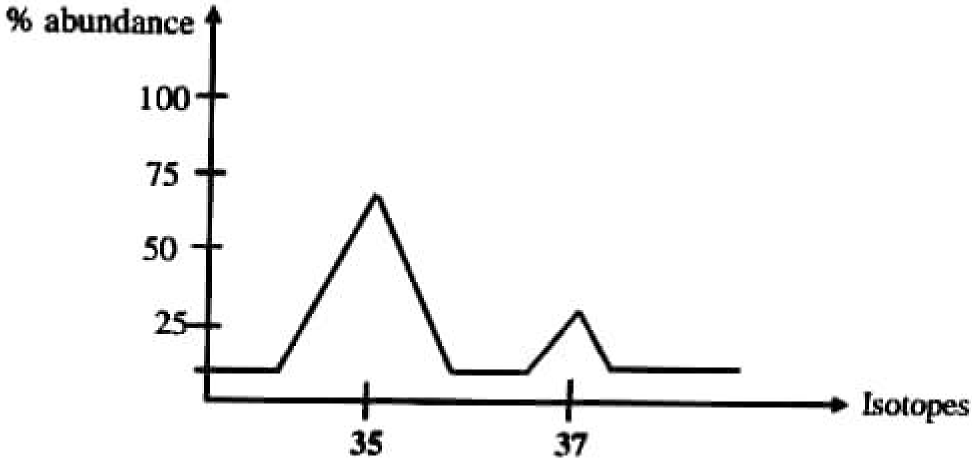
R.A.M of magnesium = $\sum Relative\ isotopic\ mass\ x\ proportion\ of\ isotope$

R.A.M = $(24x\frac{79}{100}+25x\frac{10}{100}+26x\frac{11}{100})$

R.A.M = 24.32

Therefore, relative atomic mass of 24

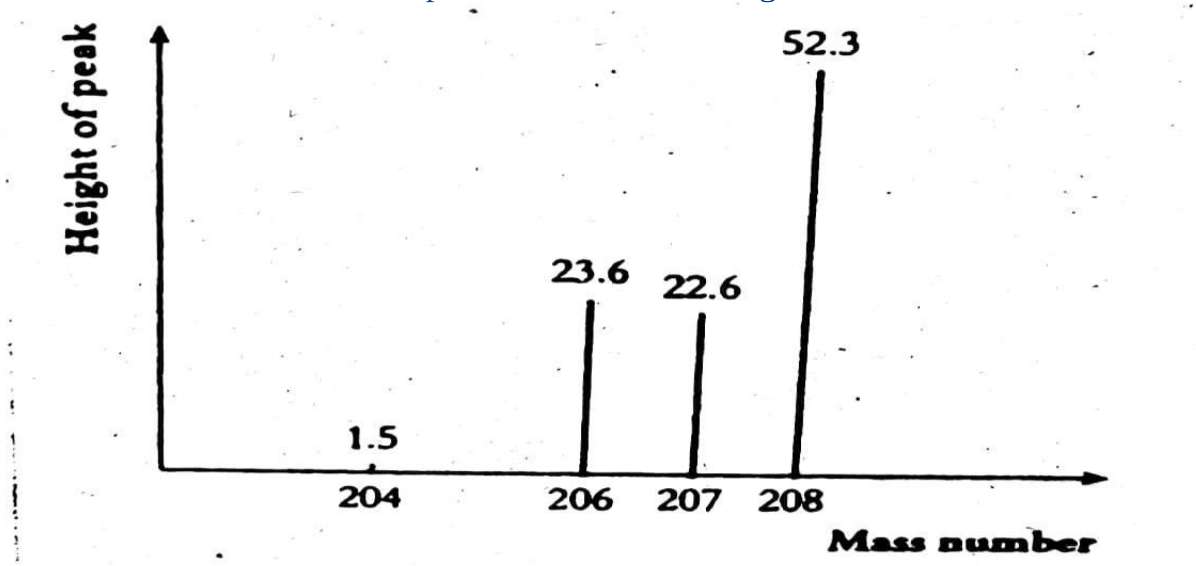
1. The figure below shows a mass spectrum for chlorine.



Task

Determine the relative atomic mass of chlorine

2. The figure below shows the mass spectrum of lead. The heights of the peaks and the mass numbers of the isotopes are shown on the figure.



Task

(a) Calculate the relative atomic mass of lead.

(b) Explain why the peaks have different heights.

3. The relative atomic mass of neon is 20.18. Naturally occurring neon has two isotopes; Neon-20 and Neon-22.

Task

(a) Calculate the percentage abundance of the isotopes.

(b) Calculate the number of neon-22 atoms in a 13.2g sample of naturally occurring neon.

4. A battery recycling plant in Uganda collects scrap lead from old batteries. Scientists suspect that some of the lead samples are not pure and may contain

unusual isotopic compositions. A mass spectrometer is used to analyze the lead sample, producing the following mass spectrum:

Mass Number	***Relative Abundance (%)***
204	*1.5*
206	*23.6*
207	*22.6*
208	*52.3*

Task

a) Calculate the relative atomic mass (RAM) of lead in the recycled sample.

b) Compare your calculated RAM with the standard atomic mass of lead (207.2). Does the sample appear to be pure?

c)Explain why different isotopes have different peak heights in the mass spectrum.

5. A water treatment plant in Uganda relies on chlorine gas (Cl_2) for purifying drinking water. Recently, concerns have arisen about the quality of the chlorine being used, which might affect the efficiency of disinfection. To investigate, scientists collect a sample of chlorine gas and analyze it using a mass spectrometer, producing the following data:

Isotope	***Mass Number (m/z)***	***Relative Abundance (%)***
Chlorine-35	*35*	*75*
Chlorine-37	*37*	*25*

Tasks:

a) Calculate the relative atomic mass (RAM) of chlorine

b) Explain why chlorine has two peaks in the mass spectrum instead of one.

c) How does mass spectrometry help ensure safe drinking water?

Sub topic 1.2: Amount of Substance, Mole Calculations

A mole is the amount of a substance which contain as many elementary units (particles, i.e., molecules, atoms, ions etc.) as there are in 12 g of carbon-12 isotope.

Molar Mass is the mass in grams of 1 mole of a substance. It is numerically equal to its relative atomic mass or its relative formula mass. e.g., One mole of carbon weighs 12 g, 1 mole of oxygen molecule weighs 32 g, and 1 mole of ammonium sulphate weighs 132 g.

Avogadro's Number,

Avogadro's number is one of the fundamental constants of chemistry. It permits one to compare the different atoms or molecules of given substances where the same number of atoms or molecules are being compared.
It also makes possible determination of how much heavier a simple molecule of one gas is than that of another, as a result the relative molecular weights of gases can be ascertained by comparing the weights of equal volumes.

1 mole of any substance contains 6.02 X 10^{23} particles.
The number of particles in any mole of a substance (6.022 x 10^{23}) is called ***Avogadro's Number,***

Calculations Using The Avogadro Number

Example 1

Calculate the number of particles in:

a) *8g of carbon (C = 12)*
b) *48g of copper (Cu = 64)*
c) *0.1 moles of sodium*
d) *0.192 moles of carbon*

Solutions

a) *1 mole of carbon weighs 12g*
1 mole of carbon contain $6.02x10^{23}$ *particles*

Therefore,
12g of carbon contains $6.02x10^{23}$ *particles*

8g of carbon will contain $\frac{8x6.02x10^{23}}{12}$ *particles*

$\underline{= 4.01x10^{23} \text{ particles}}$

b) *1 mole of copper weighs 64g*
1 mole of copper contain $6.02x10^{23}$ *particles*

Therefore,

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64g of copper contain $6.02x10^{23}$ *particles*

48g of copper contain $\frac{48x6.02x10^{23}}{64}$ *particles*

$\underline{= 4.515x10^{23}\ particles}$

c) *1 mole of sodium contain* $6.02x10^{23}$ *particles*
0.1 moles of sodium will contain $0.1x6.02x10^{23}$ *particles*

$\underline{= 6.02x10^{22}\ particles}$

d) *1 mole of carbon contain* $6.02x10^{23}$ *particles*
0.192 moles of carbon will contain $0.192x6.02x10^{23}$ *particles*

$\underline{= 1.156x10^{23}\ particles}$

Activity

1. A doctor is preparing a breathing apparatus for a patient requiring oxygen therapy. The device is filled with 40 grams of pure oxygen (O_2).
Calculate the number of molecules of oxygen in the apparatus?

2. A quarry extracts 50 grams of calcium carbonate ($CaCO_3$) for use in construction materials.
What is the number of calcium ions (Ca^{2+}) in the sample?

3. In a laboratory reaction, a scientist uses 24 grams of magnesium (Mg) to react with hydrochloric acid.
Determine the number of magnesium atoms involved in the reaction?

The 'mole' concept in calculations

Calculating Moles from Mass

The number of moles in a given mass can be calculated using:

Number of moles = $\frac{mass\ (g)}{molar\ mass\ (g/mol)}$

Examples

1. calculate the number of moles of carbon dioxide molecules present in 11 g of carbon dioxide, CO_2?

Solution

Number of moles $= \frac{mass\ of\ substance\ (g)}{molar\ mass\ /mass\ of\ 1\ mole\ of\ substance\ (g/mol)}$

$= \frac{11}{44}$

$\underline{= 0.25\ moles}$

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Alternatively

1mole of carbon dioxide molecules weighs 44g
Therefore,
44g of carbon dioxide molecule contains 1 mole

11g of carbon dioxide molecule contain $\frac{11x1}{44}$ *moles*

$= 0.25moles$

Therefore, 11g of carbon dioxide molecules contains 0.25moles

2. Find the number of moles in 80 g of magnesium oxide. (Mg = 24, O = 16).

Solution.

Molar mass of MgO = (24x1) + (16x1) = 40g/mol

40g of magnesium oxide contain 1 mole

80g of magnesium oxide will contain $\frac{80x1}{40}$ *moles*

$= 2moles$

Activity

1. *A hospital maintains emergency oxygen cylinders for patients with breathing difficulties. A technician is checking the oxygen stock and finds that one cylinder contains 50 g of oxygen gas (O_2).*
   i) Determine the number of moles of oxygen gas available in the cylinder. (O = 16)
2. *A farmer is treating acidic soil using limestone (calcium carbonate, $CaCO_3$). He collects a sample from his field and finds it weighs 10 g. To determine how much he needs for his farm, he decides to calculate the amount of calcium carbonate in moles.*
   Calculate the number of moles of calcium carbonate in the farmer's sample. (Ca = 40, C = 12, O = 16)

Calculating mass from moles

Mass in grams of substance = No. of moles **x** mass of 1 mole.

Examples

1. *What mass of magnesium (Mg=24) would contain the same number of atoms as 4 g of carbon. (C=12)*

Solution

Number of moles of carbon $= \frac{mass\ in\ grams}{mass\ of\ 1\ mole}$

$= \frac{4}{12}$ *moles*

$= 0.333moles$

A mole of any element contains the same number of atoms
Hence 0.333moles of magnesium contain the same number of atoms as in 4g of carbon
Mass of 0.333moles of magnesium = number of moles x mass of 1mole
= (0.333x24) g
= 7.992g

2. A chemical reaction produces 4 moles of aluminum oxide (Al_2O_3). Calculate the mass of Al_2O_3 produced. (Al = 27, O = 16)

Solution.
Mass of 1 mole of Al_2O_3 = (27x2) + (16x3) = 102g
102g of aluminum oxide is produced by 1 mole

Therefore, mass of 4 moles of Al_2O_3 = number of moles x mass of 1 mole

= (4x102) g

= 408g

The mass of aluminum oxide produced by 4 moles is 408g.

Activity

1. *A blacksmith is forging metal tools and needs 0.5 moles of pure iron (Fe) to create a new set of hammers. Before cutting the raw material, he must determine how much iron (in grams) is required.*
   Calculate the mass of 0.5 moles of iron needed for the blacksmith's work. (Fe = 56).
2. *A medical technician is preparing a sodium chloride (NaCl) solution for use in intravenous (IV) fluids. To ensure the correct concentration, the technician measures 2.5 moles of NaCl but needs to convert this into grams for accurate weighing.*
   Calculate the mass of 2.5 moles of sodium chloride required to prepare the IV solution. (Na = 23, Cl = 35.5)

Moles and gases

Many substances exist as gases. If we want to find the number of moles of a gas, we can do this by measuring the volume rather than the mass.
Chemists have shown by experiment that 1 mole of any gas occupies a volume of approximately 24 dm3 (24 litres) at room temperature and pressure (r.t.p.). This quantity is also known as the molar gas volume, *V*m.
Therefore, it is relatively easy to convert volumes of gases into moles and moles of gases into volumes using the following relationship:

$$\text{Number of moles of a gas} = \frac{volume\ of\ the\ gas(in\ dm^3)}{total\ volume\ at\ r.t.p\ (24dm^3)}$$

Or volume of gas (in dm^3 at r.t.p) = number of moles x total volume at r.t.p($24dm^3$)

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Example

1. Calculate the number of moles of ammonia gas, NH_3, in a volume of $72dm^3$ of the gas measured at r.t.p.

Solution

Volume of gas = $72dm^3$
Total volume at r.t.p = $24dm^3$
Number of moles of ammonia = $\frac{72}{24}$

$= 3$

Activity

1. *Hydrogen gas is produced in a reaction, and **5 moles** of hydrogen gas are collected. What volume does this hydrogen gas occupy at r.t.p.?*
2. *A reaction produces **0.75 moles** of nitrogen gas (N_2). What volume will this gas occupy at room temperature and pressure?*
3. *A sample of carbon dioxide gas has a volume of **48 dm³** at room temperature and pressure. How many moles of carbon dioxide are present?*

Moles and chemical equations

When we write a balanced chemical equation, we are indicating the numbers of moles of reactants and products involved in the chemical reaction. Consider the reaction between magnesium and oxygen.

magnesium + oxygen → magnesium oxide

$2Mg(s) + O2(g) \rightarrow 2MgO(s)$

This shows that 2 moles of magnesium react with 1 mole of oxygen to give 2 moles of magnesium oxide. Using the ideas of moles and masses we can use this information to calculate the quantities of the different chemicals involved.

$2Mg(s)$ +	$O_2(g)$	$\rightarrow$	$2MgO(s)$
2 moles	1 mole		2 moles
2×24	$1 \times (16 \times 2)$		$2 \times (24 + 16)$
$= 48$ g	$= 32$ g		$= 80$ g

Example

1. Calculate the amount of lime produced when 10g of limestone are heated (C = 12; O = 16; Ca = 40)

Solution

$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$
Molar mass of $CaCO_3 = (40x1 + 12x1 + 16x3) = 100g$
Molar mass $CaO = (40x1+16x1) = 56g$

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From equation

100g of calcium carbonate decompose to produce 56g of lime (calcium oxide)

10g of calcium carbonate will decompose to produce ($\frac{10x56}{100}$)g of lime (CaO)

The amount of lime produced is 5.6g

Activity

1. *During an outdoor camping event, a group of students used a magnesium strip to create a bright light source. They burned* ***2.4 g*** *of magnesium in an open area where oxygen was in excess. After the reaction, they collected the white powder formed. Calculate the mass of the white powder (magnesium oxide) produced when the magnesium strip was completely burned.*
2. *A bakery owner in Kampala noticed that when baking soda (sodium hydrogen carbonate) is heated, it releases a gas that makes bread rise. To improve his recipe, he decided to study how much carbon dioxide is released when 42 g of baking soda is heated completely.*
   Determine the mass of the new compound (sodium carbonate) formed after heating
3. *In a local water treatment plant, industrial wastewater containing 4.0 g of copper (II) sulfate was found in a sample. To remove the toxic copper ions, hydrogen sulfide gas was bubbled through the solution, forming an insoluble black precipitate of copper (II) sulfide that could be filtered out. Calculate the mass of copper (II) sulfide that was precipitated during the purification process*

Calculation of Percentage composition:

Example

1. Calculate the percentage composition by mass of magnesium oxide, (Mg = 24, O = 16).

Solution

formula of the compound: MgO

formula mass of the compound: 24 + 16 = 40

Expressing each atomic mass as a percentage of the formula mass.

Percentage of Magnesium $= \frac{24}{40}x100 = 60\%$

Percentage of oxygen $= \frac{16}{40}x100 = 40\%$

Activity

1. *A soap-making business in Uganda uses hydrated sodium carbonate ($Na_2CO_3 \cdot 10H_2O$), commonly known as washing soda, in its production. However, the workers notice that when left in open containers, the washing soda seems to lose weight over time. They suspect it might be losing water of crystallization.*
   Calculate the percentage of water of crystallization in hydrated sodium carbonate ($Na_2CO_3 \cdot 10H_2O$). (Na = 23, C = 12, H = 1, O = 16)
2. *A matchstick production company burns sulfur in excess oxygen to produce sulfur dioxide (SO_2), a key chemical in the manufacturing process. One day,*

the factory produces 640 g of sulfur dioxide and the chemist needs to determine how much sulfur was used.
Calculate the mass of sulfur that was burned to produce 640 g of sulfur dioxide. (S = 32, O = 16)

3. *A scrap metal dealer collects copper from old electrical wires and heats it in excess oxygen to form copper (II) oxide (CuO), which can then be processed into pure copper. If the process produces 159 g of copper (II) oxide, the dealer wants to know how much pure copper was originally present.*
Calculate the mass of copper required to produce 159 g of copper (II) oxide when heated in excess oxygen. (Cu = 64, O = 16)

SUB-TOPIC 1.3: Chemical Formulae and Chemical Equations

Every element has a symbol. A symbol is a letter or two letters which stand for one atom of the element, formulae are written for compounds.
A **chemical formula** represents a substance using symbols for its constituent elements and numerical subscripts to indicate atomic ratios.

Chemical formulae of some compounds

Compound	Chemical formula
Sodium carbonate	Na_2CO_3
Sodium hydrogen carbonate	$NaHCO_3$
Water	H_2O
Calcium sulphate	$CaSO_4$
Carbon dioxide	CO_2

Activity

Write the chemical formulae of the following compounds

a) *Sodium hydroxide*
b) *Ammonia*
c) *Hydrochloric acid*
d) *Calcium chloride*
e) *Iron III sulphate*

Chemical equations

Chemical reactions involve the transformation of reactants into products through the breaking and forming of chemical bonds. To represent these transformations clearly, **chemical equations** are used. These equations rely on **chemical formulae**, which depict the composition of substances in terms of element symbols and subscripts.

Steps in writing equation

1. Establish the formulae of all the substances taking part in the reaction, both as reactants and products
2. Write down the formulae of the reactants and the products in some agreed manner. Conventionally the reactants are placed on the left-hand side of the page and the products on the right-hand side. An arrow from the left to right indicates the reaction proceeds from reactants to products as written

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3. Ensure that the number of atoms on both side of the reaction are equal and write correct states. These states are usually written using state symbols which are place after the formula of the substance in the equation, and are usually enclosed in parentheses. The table below gives the principle state symbols.

State of the substance	State symbol
Solid	S
Liquid	L
Gas	G
Aqueous solution	aq

Thus, the reaction between sodium chloride solution ad lead nitrate solution to give a precipitate of lead (II) chloride maybe written:
$2NaCl(aq) + Pb(N0_3)_2(aq) \longrightarrow PbCl_2(s) + 2NaNO_3(aq)$

Note: you can start by writing words equation and replacing names with formulae when writing chemical equation for the reaction.

Examples

a) *Calcium carbonate decomposes to give calcium oxide and carbon dioxide.*
1. *Calcium carbonate $\longrightarrow$ calcium oxide + carbon dioxide*
2. $CaCO_3 \longrightarrow CaO + CO_2$
3. $CaCO_3(s) \longrightarrow CaO(s) + CO_2(g)$
4. *On the left-hand side, you have 1 atom of calcium, 1 atom of carbon and 3 atoms of oxygen combined as calcium carbonate. On the right-hand side, you have 1 atom of calcium and 1 atom of oxygen combined as calcium oxide and 1 atom of carbon, 2 atoms of oxygen combined as carbon dioxide. The two side are equal.*

Activity

1.Write equation for the following chemical reactions

a) *Magnesium reacts with sulphuric acid to give hydrogen gas and solution of magnesium sulphate*
b) *Hydrogen reacts with oxygen to form water*
c) *Sodium carbonate reacts with dilute hydrochloric acid to give carbon dioxide and solution of sodium chloride*
d) *Carbon burn in excess air forming carbon dioxide gas*
e) *Silver nitrate reacts with sodium bromide forming silver bromide and sodium nitrate.*

Relationship between empirical formulae and molecular formulae

Empirical formula of a compound is the simplest formula which represents its composition. It shows the elements present and the ratio of the amounts of elements present.

The molecular formula (M.F) of a compound is the one which shows the exact number of each kind of each atom present in a molecule of a compound.
The molecular formula is always a simple multiple of the empirical formula.
i.e., M.F = n(E.F) where n is a whole number e.g., E.F of glucose is CH_2O, whereas its M.F. is $C_6H_{12}O_6$ which $(CH_2O)_6$.

Experiment to determine the empirical formula of magnesium oxide

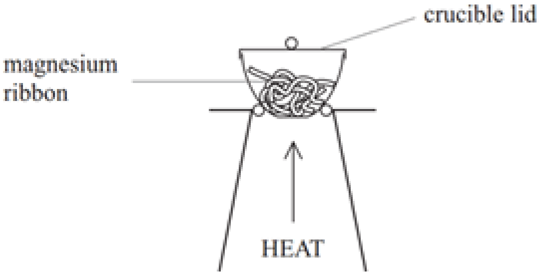
Materials and Apparatus

- *Magnesium ribbon (cleaned)*
- *Crucible with a lid*
- *Digital balance (accurate to 0.01 g)*
- *Bunsen burner*
- *Tongs*
- *Tripod stand*

Procedure

A known mass of magnesium ribbon is heated in a crucible strongly
After the magnesium had all burnt, it is allowed to cool, the crucible with its content/products (white solid) is reweighed.

Setup



Results

Mass of Mg = mass of crucible + mass of Mg – mass of crucible alone
Mass of MgO = mass of crucible + mass of MgO – mass of crucible alone
Mass of O_2 reacted = mass of MgO – mass of Mg

Treatment of results

Element	Magnesium (Mg)	Oxygen (O)
Mass in gram		
Moles		
Mole ratio		
Number of atoms		

Therefore, empirical formula will be based on the number of atoms of the elements

Calculations involving empirical formulae and molecular formulae

Example

1. *Determine the formula of a lead compound, given that 4.14 g of lead combines with 0.64 g of sulphur and 1.28 g of oxygen.*

Solution

Elements	Lead (Pb)	Sulphur (S)	Oxygen (O)
Mass in g	4.14	0.64	1.28
R.A.M	207	32	16
Number of moles	$\frac{4.14}{207}$ = 0.02	$\frac{0.64}{32}$ = 0.02	$\frac{1.28}{16}$ = 0.08
Ratio of moles	$\frac{0.02}{0.02}$ = 1	$\frac{0.02}{0.02}$ = 1	$\frac{0.08}{0.02}$ = 4
Number of atoms	1	1	4
Empirical formula is $PbSO_4$			

2. *The combustion of 0.92 g of copper gave 1.15 g of copper oxide. Calculate the percentage of copper and oxygen in the sample and hence its formula.*

Solution

Mass of oxygen in the oxide = (1.15 – 0.92) = 0.23g

Percentage composition of copper = $(\frac{0.92}{1.15} x100) = 80\%$

Percentage composition of oxygen = $(\frac{0.23}{1.15} x100) = 20\%$

Element	Copper (Cu)	Oxygen (O)
Percentage composition	80	20
Number of moles	$\frac{80}{64}$ = 1.25	$\frac{20}{16}$ = 1.25
Mole ratio	$\frac{1.25}{1.25}$ = 1	$\frac{1.25}{1.25}$ = 1
Number of atoms	1	1

Empirical formula is CuO

Activity

3. *3.22 g of hydrated sodium sulphate, Na2SO4.nH2O were heated till all the water of crystallization was driven off. The anhydrous salt left had a mass of 1.42 g. Determine the formula of the hydrated salt.*
4. *An organic compound was found to contain 12.8% carbon, 2.1% Hydrogen and 85.1% bromine. A rough estimate gave its relative molecular mass between 150 and 200. Determine its molecular formula. Explain your answer.*

3. Calculate the empirical formula of an organic compound containing 92.3% carbon and 7.7% hydrogen by mass. The Molecular mass of the organic compound is 78. What is its molecular formula? (H = 1; C = 12)

Solution

Elements	C	H
Percentage by mass	92.3	7.7
Masses in 100g	92.3g	7.7g
Number of moles	$\frac{92.3}{12} = 7.7$	$\frac{7.7}{1} = 7.7$
Mole ratio	$\frac{7.7}{7.7} = 1$	$\frac{7.7}{7.7} = 1$
Number of atoms	1	1

Therefore, empirical formula is CH

From $(Emf)_n$ = *molecular mass*

$(CH)_n = 78$

$(12x1 + 1x1)_n = 78$

$(12 + 1)_n = 78$

$13n = 78$

$n = \frac{78}{13}$

$n = 6$

Therefore, from $(Emf)_n$,

$(CH)_6$

Therefore, molecular formula of the organic compound is C_6H_6

Activity

1. *A fuel research company in Uganda is testing a new hydrocarbon-based fuel for vehicles. A sample of the fuel was analyzed and found to contain 80% carbon and 20% hydrogen by mass. To determine how efficiently it burns, the researchers need to find its empirical and molecular formula.*
   a) *Determine the empirical formula of the hydrocarbon fuel.*
   b) *Given that its molecular mass is* ***30****, find its molecular formula.*
2. *Combustion of 5.4 g of a compound containing C, H and O only gave 7.92 g of CO2 and 3.24 g of H2O. Calculate the empirical formula of the compound? The compound has a R.M.M of 180, determine its molecular formula.*
3. *A factory storing iron rods in a humid environment noticed that when exposed to chlorine gas, the iron reacted and formed an iron chloride compound. A laboratory test revealed that 2.8 g of iron reacted with 5.3 g of chlorine to form this chloride.*
   a) *Determine the simplest formula (empirical formula) of the iron chloride.*
   b) *Given that the molecular mass of the chloride is 325, find its molecular formula.*

SUB-TOPIC 1.4: Solutions and Concentration, Calculations Involving Gas Volumes

Definition of common terms

A solution is a homogeneous mixture of two or more substances, where one substance (the solute) is dissolved in another (the solvent). The resulting mixture has uniform composition throughout.

The concentration of a substance is the number of moles or the mass of a solute dissolved or contained in a known volume of solution.
Usually, the concentration of a solution is expressed in either number of grams or number of moles of solute per litre of solution.

Concentration in moll^{-1}/moldm^{-3} $= \frac{\text{amount of solute in moles}}{\text{volume of solution in dm}^3}$

Or concentration in gl^{-1} = number of moles in 1l/1dm^3 x molar mass/formula mass

The **Molarity** of solution is the number of moles of the solute contained in 1 litre, 1000 ml or 1000 cm3 of solution.

Molarity of a solution $= \frac{mass\ of\ solute\ dissolved\ in\ 1\ mole}{molar\ mass/formula\ mass}$

A **molar solution** is a solution containing 1 mole of a solute/substance per litre.

A standard solution is a solution whose concentration is accurately known i.e. one that contains a known amount of solute in a known volume of solution.

Preparation of a standard solution

A standard solution is prepared from a substance called a primary standard.
A primary standard is a substance that is analytically pure and chemically stable such that a known mass of it when weighed is the exact mass that dissolves in water to form a standard solution.

A primary standard has the following characteristics;

1. Readily soluble in water at room temperature such that all the weighed mass goes into solution.
2. It should have a fairly high molar mass such that weighing errors are minimized.
3. It should be obtainable in a pure state such that the quantity weighed indicates the actual mass present in the standard solution prepared.
4. It should not be deliquescent, efflorescent or hygroscopic such that the mass weighed is exactly that of the pure sample
5. It must be stable at ordinary temperatures (should not decompose easily) such that its chemical nature is not altered.
6. It should be able to undergo stoichiometric and instantaneous reactions such that titration errors are minimized.

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Preparation of standard solutions by weighing and dissolving in water

Preparing a standard solution of anhydrous sodium carbonate (for example 250cm3 of 0.1M sodium carbonate solution)

a) Calculate the mass of anhydrous sodium carbonate needed to make the standard solution.

250cm3 of 0.1M sodium carbonate solution

Number of moles of $Na_2CO_3 = (\frac{250x0.1}{1000}) = 0.025moles$

Molar mass of $Na_2CO_3 = (23x2) +(12x1) +(16x3) = 106g$

1 mole of Na_2CO_3 weigh 106g

0.025 moles of Na_2CO_3 weigh $(\frac{0.025x106}{1})g = 2.65g$

2.65g of Na_2CO_3

b) A clean container is weighed, and record its mass. Use a clean spatula and add pure anhydrous sodium carbonate to the weighing container until the mass is the total mass of the weighing container and that of anhydrous sodium carbonate required.

c) Transfer the weighed mass of anhydrous sodium carbonate carefully into a clean beaker. Using a wash bottle of distilled water, add a small volume of water so that all the washings run into the beaker. Add about 100 cm^3 of distilled water and stir with a glass rod until all the solid has dissolved.

d) Pour all the solution carefully through a filter funnel into a volumetric flask. Wash all the solution out of the beaker and off the glass rod to the volumetric flask.

e) Add distilled more distilled water carefully into the solution in a volumetric flask until the level of the solution is on the mark.

f) Label the solution as required. The solution formed will be 0.1M Na_2CO_3

Preparation of standard solution by diluting a concentrated solution

Example

Preparation of 0.1M sodium hydroxide solution from a stock solution which is 2M

Calculate the volume of the stock solution needed to make the standard solution of 0.1M NaOH

2moles of concentrated solution are in 1000cm^3 of the solution

0.1 moles of concentrated solution are in $(\frac{0.1x1000}{2})$ cm^3 of the solution

$= 50cm^3$ of the concentrated solution

Alternatively

From

Molarity of concentrated solution M_1 x its volume V_1 = molarity of dilute M_2 x its volume V_2

$$M_1 x V_1 = M_2 V_2$$

$$2xV_1 = 0.1 \text{ x } 1000$$

$$V_1 = \frac{0.1x1000}{2}$$

$$V_1 = 50cm^3$$

Procedures

- *Measure 50cm3 of concentrated solution of sodium hydroxide into a 1000cm3 measuring cylinder*
- *Transfer the solution into a volumetric flask*
- *Make up to the mark with more distilled water. Stir*
- *The solution formed is 0.1M sodium hydroxide solution*

Preparation of 0.1M H_2SO_4 from a container of concentrated sulphuric acid having the following specifications

- Molecular weight = 98.08g
- Minimum assay = 98%
- Specific gravity(20^{o}c)/weight per mililitre = 1.84g

Calculate the volume of the stock solution needed to make the standard solution of 0.1M H_2SO_4

From the specifications

1cm^3 of concentrated H_2SO_4 solution weighs 1.84g

1000cm^3 of concentrated H_2SO_4 solution weigh $(\frac{1000x1.84}{1})g$

$$= 1840gdm^3$$

But 98g of H_2SO_4 makes 1 mole

1840g of H_2SO_4 makes $(\frac{1840x1}{98})$moles

$$= 18.78moldm^{-3}$$

From

$$M_1V_1 = M_2V_2$$
$$18.78V_1 = 0.1x1000$$

$$V_1 = (\frac{0.1x1000}{18.78})$$

$$V_1 = 5.32cm^3$$

Procedures

- *Withdraw 5.32cm^3 of the concentrated acid solution using a measuring cylinder*
- *Put some distilled water in a 1 litre measuring cylinder for example 500cm^3 of distilled water*
- *Add the 5.32cm^3 concentrated acid solution to the distilled water in a measuring cylinder slowly.*
- *Make up to the mark with more distilled water and the label the solution*
- *The solution formed is a 0.1M H_2SO_4 acid*

Activity

1. *Viola, a newly employed laboratory technician at a secondary school, has been tasked by the chemistry teacher to prepare 250 cm³ of a 0.2 M sodium hydroxide (NaOH) solution using sodium hydroxide pellets. However, since this is her first-time handling solution preparation, she is unsure of the exact steps to follow.*
   Carry out an experiment to prepare 250cm3 of 0.2M sodium hydroxide solution, use you experiment to explain to viola the necessary steps to follow.
2. *A bottle of concentrated hydrochloric acid has the following specifications*
   - *Molecular weight = 36.5g*
   - *Minimum assay = 36%*
   - *Density/specific gravity/weight per ml = 1.18g*

   Using the information above, prepare 250cm3 of 0.2M hydrochloric acid.

Calculations involving solutions

Example

1. 25cm^3 of sodium hydroxide solution were neutralized by 20.50cm^3 of 0.115M hydrochloric acid. Calculate the molarity and the concentration of sodium hydroxide in gl-1 (Na = 23, O = 16, H = 1)

Solution

Moles of hydrochloric acid that reacted = $\left(\frac{0.115x20.50}{1000}\right)$ *moles*

= *0.0023575moles*

Equation of the reaction

$NaOH_{(aq)} + HCl_{(aq)} \longrightarrow NaCl_{(aq)} + H_2O_{(l)}$

Mole ratio of HCl : NaOH = 1:1

Reacted moles of NaOH = 0.0023575 moles

Concentration of NaOH in moll1- = $\left(\frac{0.0023575x1000}{25}\right)$ *moles*

Molarity of NaOH = 0.0943moll^{-1}

Concentration in gl^{-1} = molarity x molar mass
But molar mass of NaOH = (234x1) +(16x1) +(1x1) = 40g
Therefore, concentration in gl-1 = 0.0943x40
= $3.772gl^{-1}$

Activity

1. *$25cm^3$ of 0.05M sulphuric acid neutralizes $35cm^3$ of potassium hydroxide solution. What is the molarity and concentration in gl^{-1} of potassium hydroxide?*
2. *15g of anhydrous sodium carbonate containing some sodium chloride as impurity were made up to $250cm^3$ of solution. $25cm^3$ of this solution required $24.5cm^3$ of 1M hydrochloric acid for complete neutralization. Calculate the percentage of sodium chloride in the solid. (Na = 23, C = 12, O = 16)*

Using standard solution to standardize other solutions.

Standardizing *refers to the process of determining the exact concentration (molarity) of a solution by reacting it with a solution of a known concentration (a primary standard). This process is commonly used in titration experiments to ensure accuracy in quantitative chemical analysis.*

1. Standardization of hydrochloric acid using Borax

5.7g of disodium tetraborate decahydrate (Borax) were dissolved in 100 cm^3 distilled water in a $250cm^3$ volumetric flask, shaken to dissolve and the mixture made up to the mark with distilled water. 25.0 cm^3 of the resultant solution is pipetted into a clean conical flask and titrated against dilute hydrochloric acid solution which was approximately 0.1M using phenolphthalein indicator. 18.0 cm^3 of the acid was required for complete neutralisation. Calculate the molarity of the hydrochloric acid solution used. (Na = 23, B = 11, O =16, H = 1)

Solution

Moles of $Na_2B_4O_7.10H_2O$ in $250cm^3$ of the solution
From
Molar mass of $Na_2B_4O_7.10H_2O$ = (23x2) +(11x4) +(16x3) +(10x18) = 382g
Therefore,
382g of Borax contains 1 mole

5.7g of Borax contain $(\frac{5.7x1}{382})$ *moles*

= 0.0149moles
Therefore, $250cm^3$ of the solution contain 0.0149 moles of borax
Reacted moles of Borax
$250cm^3$ of the solution contain 0.0149 moles of borax

$25cm^3$ of the solution contain $(\frac{25x0.0149}{250})$ *moles*

Organized by JASON Head of department science cornerstone s s
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Reacted moles of Borax is 0.00149moles

Equation of the reaction

$B_4O_7^{2-}{}_{(aq)} + 2H^+{}_{(aq)} + 5H_2O_{(l)} \longrightarrow 4H_3BO_{3(aq)}$

Mole ratio of Borax: acid is 1:1

Reacted moles of acid = reacted moles of Borax = 0.00149 moles

Molarity of the acid

18.0cm³ of the acid solution contain 0.00149 moles

1000cm³ of the acid solution contain $\left(\frac{1000x0.00149}{18.0}\right)$ *moles*

Molarity of the acid is 0.08M

Activity

1.50g of anhydrous sodium carbonate were dissolved in 100cm³ distilled water in a 250cm³ volumetric flask, shaken to dissolve and the mixture made up to the mark with distilled water. 20.0cm³ of the resultant solution is pipetted into a clean conical flask and titrated against dilute hydrochloric acid solution using methyl orange indicator. 27.50cm³ of the acid was required for complete neutralisation. Calculate the molarity of the hydrochloric acid solution used.

PRACTICAL APPLICATION

Standardisation of hydrochloric using anhydrous sodium carbonate.

You are provided with the following:

***FA1** which is 60cm³ of hydrochloric acid solution*

*Solid **Y** which is anhydrous sodium carbonate*

Methyl orange indicator

*You are required to determine the molarity of **FA1** using a standard solution of anhydrous sodium carbonate.*

Procedure:

*Accurately measure 40cm³ of **FA1** and add exactly 60cm³ of distilled water. Label the resultant solution **FA3.***

*Weigh accurately 1.5g of solid **Y** into a clean 250cm³ volumetric flask and add about 100cm³ of distilled water. Shake well to dissolve. Add more distilled water to fill up to the mark and label the resultant solution **FA2***

*Pipette 25.0 cm³ (or 20.0 cm³) of **FA2** into a clean conical flask. Add 2 or 3 drops of methyl orange indicator.*

*Titrate the resultant mixture against **FA3** from the burette until the endpoint is reached.*

Repeat the titration until you obtain consistent results. Record your results in the table below.

Mass of Y and weighing bottle …………………… g

Mass of empty weighing bottle …………………… g

Mass of Y used ……………………………………….. g

Volume of pipette used ……………………………. cm³

Final burette reading (cm^3)			
Initial burette reading (cm^3)			
Volume of FA3 used (cm^3)			

Record the titre values used to calculate the average volume of **FA3** used.
……………………………………………………………………………cm^3

Average volume of FA3 used ……………………………………………………cm^3

Calculate the molarity of hydrochloric acid in FA1

Activity

You are provided with:
***FA1** which is approximately 0.1M hydrochloric acid solution*
*Solid **Q** which is hydrated disodium tetraborate (Borax)*
Phenolphthalein indicator
*You are required to standardise hydrochloric acid in **FA1** using a standard solution of Borax.*

Procedure:

*Weigh accurately 4.7g of solid **Q** into a clean 250cm^3 volumetric flask and add about 100 cm3 of distilled water. Shake well to dissolve. Add more distilled water to fill up to the mark and label the resultant solution **FA3***
*Pipette 25.0cm^3 (or 20.0cm^3) of **FA3** into a clean conical flask. Add 2 or 3 drops of phenolphthalein indicator and shake the flask carefully.*
*Titrate the solution carefully against **FA1** from the burette until the endpoint is reached.*
Repeat the titration until you obtain consistent results. Record your results in the table below.

Results:

*Mass of **Q** and weighing bottle…………………………………………………………………g*
Mass of empty weighing bottle…………………………………………………………………g
*Mass of **Q** used……………………………………………………………………………………g*
Volume of pipette used…………………………………………………………………………cm^3

Final burette reading (cm^3)			
Initial burette reading (cm^3)			
Volume of FA3 used (cm^3)			

Record the titre values used to calculate the average volume of **FA1** used.
………………………………………………………………………………………………cm^3

Average volume of **FA1** used……………………………………………………………… cm^3

Calculate the concentration of hydrochloric acid in **FA1** per litre of solution.

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Gas laws

Gases are a fundamental state of matter characterized by their ability to expand and fill any container, their compressibility, and the large distances between their particles. Unlike solids and liquids, gases do not have a fixed shape or volume, making their behavior highly dependent on external conditions such as pressure, temperature, and volume. To understand and predict how gases behave under different conditions, scientists have formulated a set of mathematical relationships known as the **gas laws**.

Boyle's law

Boyle's Law states that at constant temperature, the pressure of a fixed mass of a gas is inversely proportional to the volume

Mathematically

$$PV = a\ constant(T)$$

Charles law

Charles' Law states that at constant pressure, volume of a given mass of a gas is directly proportional to its absolute temperature:

Mathematically

$$\frac{V}{T} = a\ constant(P)$$

Gay-Lussac's Law

When gases react together, they do so in volumes which are related to each other in simple ratio and to the product if gaseous, provided all measurements are made at the same temperature and pressure.

$$1\ volume\ of\ H_2 + 1\ volume\ of\ Cl_2 \longrightarrow 2\ volumes\ of\ HCl$$
$$2\ volumes\ of\ H_2 + 1\ volume\ of\ O_2 \longrightarrow liquid\ water$$
$$2\ volumes\ of\ H_2 + 1\ volume\ of\ O_2 \longrightarrow 2\ volumes\ of\ steam$$
$$1\ volume\ of\ N_2 + 3\ volumes\ of\ H_2 \longrightarrow 2\ volumes\ of\ ammonia$$

The above three laws are combined, we get a simple equation representing the relationship between pressure, volume and absolute temperature of a fixed amount(mole) of a gas
This equation is

$$\frac{PV}{T} = a\ constant$$

This is often written as

$$\frac{P1V1}{T1} = \frac{P2V2}{T2}$$

Example

Timothy collected 60cm^3 of nitrogen gas at 60^{o}c and $1.05x10^5 Nm^{-2}$. Calculate the volume of the gas at s.t.p

Solution

Experimental conditions are

$P_1 = 1.05x10^5 Nm^{-2}$
$T_1 = 273+60 = 333K$
$V_1 = 60cm^3$

Standard conditions are

$P_2 = 1.01x10^5 Nm^{-2}$
$T_2 = 273K$

From $\frac{P1V1}{T1} = \frac{P2V2}{T2}$

$$V_2 = \frac{1.05x10^5x60x273}{1.01x10^5x333} cm^3$$

$= 51cm^3$

The volume of the gas at s.t.p is 51cm^3

Activity

1. *Calculate the volume of oxygen at 12^{o}c and 745mmHg pressure which could be obtained by heating 5g of potassium chlorate (V). (K = 39, Cl = 35.5, O = 16. Molar volume of gases at s.t.p is 22.4dm^3)*
2. *Calculate the volume of hydrogen sulphide at 14oc and 770mmHg pressure which will react with 10g of lead (II) nitrate. (Pb = 207, N = 14, O = 16. Molar volume of gases is 22.4dm^3 at s.t.p)*

Avogadro's law

The Italian scientist Avogadro, in 1811 suggested that in gaseous elements like chlorine and hydrogen, the atoms joined to form larger particles containing two or more atoms.

He called the larger particles **molecules**. *A molecule is therefore defined as a group of like or unlike atoms, chemically combined together*.

Avogadro came up with a hypothesis that then became a theory to now a law which states that; ***Equal volumes of all gases measured at the same conditions of temperature and pressure contain the same number of molecules***.

If hydrogen and chlorine were to combine therefore;

$$1\ volume\ of\ H_2 + 1\ volume\ of\ Cl_2 \longrightarrow 2\ volumes\ of\ HCl$$

can be interpreted as;

$$1\ molecule\ of\ H_2 + 1\ molecule\ of\ Cl_2 \longrightarrow 2\ molecules\ of\ HCl$$
$$n\ molecules\ of\ H_2 + n\ molecules\ of\ Cl_2 \longrightarrow 2n\ molecules\ of\ HCl$$
$$equation; \quad H_2 + Cl_2 \longrightarrow 2HCl$$

It follows from Avogadro's hypothesis that if equal volumes of gases contain equal numbers of molecules, then the volume occupied by one mole of molecules of gases must be the same for all gases. It is called **molar gas volume**.
The value is

$$22.4\ dm^3 or\ litres\ or\ 22400cm^3\ at\ stp(0°C\ and\ 1\ atm)\ \mathbf{or}\ 24\ dm^3 or\ litres\ or\ 24000cm^3\ at\ room\ temperature\ (20°C\ and\ 1\ atm)$$

It also therefore follows that ***the total number of moles of a gas is directly proportional to the volume occupied by a gas at constant temperature***.

$$\boldsymbol{V} \propto \boldsymbol{n}\ (n = number\ of\ moles\ of\ a\ gas)$$
$$\boldsymbol{V} = \boldsymbol{kn}$$
$$\frac{\boldsymbol{V}}{\boldsymbol{n}} = \boldsymbol{k};\ \ where\ V\ is\ volume\ of\ a\ gas$$
$$k\ is\ a\ constant.$$
$$\frac{\boldsymbol{V_1}}{\boldsymbol{n_1}} = \frac{\boldsymbol{V_2}}{\boldsymbol{n_2}}$$

However, Avogadro's law cannot be true at all pressures. It only holds if the gases behave as perfect gases, that is, at very low pressures.

Example

1. *One mole of helium gas fills up an empty balloon to a volume of 1.5 litres. What would be the volume of the balloon if an additional 2.5 moles of helium gas are added? (Assume that the temperature and the pressure are kept constant)*

Solution

V_1 = 1.5litres
V_2 = ??
n_1 = 1 mole
n_2 = 1+2.5 = 3.5 moles

from

$$\frac{V1}{n1} = \frac{V2}{n2} \qquad V2 = \frac{V1n2}{n1} = \frac{1.5x3.5}{1} = 5.25litres$$

The volume of balloon is 5.25litres

2. *50g of nitrogen gas are contained in a 3000cm^3 container. The gas exerts a pressure of 3 atmospheres on the container. If the pressure and temperature are kept constant, calculate the mass of nitrogen that can be added the container until the volume reaches 5000cm^3.*

Solution

Molar mass of n_2 = 14x2 = 28g

Number of moles of n_2 = $\frac{50}{28}$ = 1.7857moles

V_1 = 3000cm3
V_2 = 5000cm3
n_1= 1.7857 moles
n_2 = ??

From

$$\frac{V1}{n1} = \frac{V2}{n2} \quad n2 = \frac{V2xn1}{V1} = \frac{5000x1.7857}{3000} = 2.9762 \text{ moles}$$

Extra moles of gas added = (2.9762 – 1.7857) = 1.1905moles

Extra masses of gas added = (28x1.1905) = 33.334g.

Activity

1. *If 3.0 L of oxygen gas contains 0.12 moles, how many moles of oxygen would be present in 7.5 L of gas under the same conditions?*
2. *A gas cylinder contains 6.0 moles of carbon dioxide in a volume of 18.0 L. How much volume would the gas occupy if only 4.0 moles remain, assuming temperature and pressure remain unchanged?*

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