DEPARTMENT OF EDUCATION

GRADE 12

CHEMISTRY

MODULE 1

MASSES, MOLES AND CONCENTRATION

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Writer
Evelyn Dimacale

Content Editors
Science Department
Subject Review Committee

Language Editor
Dr. Steven Winduo

Course Format Editor
Joydee Mabbagu

GRADE 12

CHEMISTRY

MODULE 1

MASSES, MOLES AND CONCENTRATION

IN THIS MODULE YOU WILL LEARN ABOUT:

12.1.1: ISOTOPES
12.1.2: RELATIVE FORMULA MASS AND PERCENTAGE COMPOSITION
12.1.3: MOLES
12.1.4: EMPERICAL AND MOLECULAR FORMULAS
12.1.5: STOICHIOMETRY
12.1.6: SOLUTIONS

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MR. DEMAS TONGOGO
Principal-FODE

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SECRETARY'S MESSAGE

Achieving a better future by individual students and their families, communities or the nation as a whole, depends on the kind of curriculum and the way it is delivered.

This course is part and parcel of the new reformed curriculum. The learning outcomes are student-centred with demonstrations and activities that can be assessed.

It maintains the rationale, goals, aims and principles of the national curriculum and identifies the knowledge, skills, attitudes and values that students should achieve.

This is a provision by Flexible, Open and Distance Education as an alternative pathway of formal education.

The course promotes Papua New Guinea values and beliefs which are found in our Constitution and Government Policies. It is developed in line with the National Education Plans and addresses an increase in the number of school leavers as a result of lack of access to secondary and higher educational institutions.

Flexible, Open and Distance Education curriculum is guided by the Department of Education's Mission which is fivefold:

- to facilitate and promote the integral development of every individual
- to develop and encourage an education system that satisfies the requirements of Papua New Guinea and its people
- to establish, preserve and improve standards of education throughout Papua New Guinea
- to make the benefits of such education available as widely as possible to all of the people
- to make the education accessible to the poor and physically, mentally and socially handicapped as well as to those who are educationally disadvantaged.

The college is enhanced through this course to provide alternative and comparable pathways for students and adults to complete their education through a one system, two pathways and same outcomes.

It is our vision that Papua New Guineans' harness all appropriate and affordable technologies to pursue this program.

I commend all the teachers, curriculum writers and instructional designers who have contributed towards the development of this course.

DR. UKE KOMBRA, PhD
Acting Secretary for Education.

MODULE 1: MASSES, MOLES AND CONCENTRATION

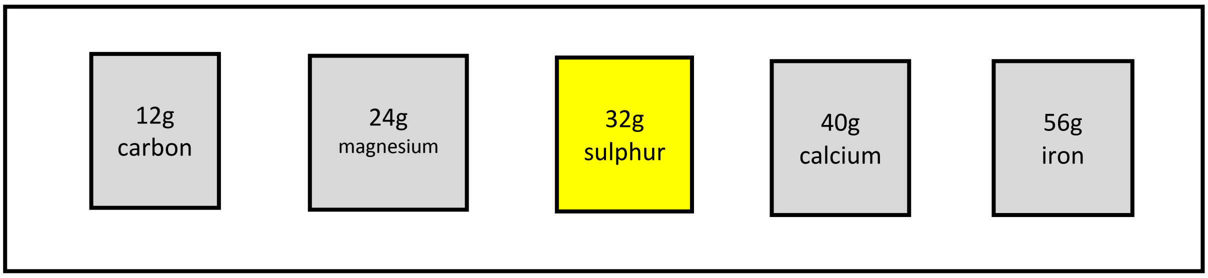
INTRODUCTION

Have you ever tried counting the number of rice particles in a bucket of rice? It is difficult to do so because rice particles are very small and numerous. Chemists face a similar problem when they try to count atoms. Atoms are too small to be counted one at a time. Because they are so small, it is difficult to measure the mass of each atom that you need to take a very large number of them to get a gram of an element or compound.

How many atoms are there in 1g of hydrogen or 12g of carbon? Scientists in the 18^{th} century were very interested to know the answers and various hypotheses were put forward. The scientists realised that there was a number, $\mathbf{N_A}$, which would be universal constant for the number of atoms in 1g of hydrogen, or 12g of carbon, or 23g of sodium. They called this number **'Avogadro's Number' ($\mathbf{N_A}$)** in honour of Amadeo Avogadro, and the number is $\mathbf{6.02 \times 10^{23}}$ which is equal to one mole of any atoms, ions or molecules.

Do you know how large 1 mole or 6.02×10^{23} is? If you have one mole of people, they would cover the whole surface of earth and oceans and the column of people, one standing on top of the other, and would stretch beyond the moon. If you have one mole of footballs packed closely together, they would occupy the same volume as the earth.

Take the relative atomic mass in grams of any element like the one shown below. They are all one mole each and contain the same number of atoms, 6.02×10^{23}.



All these masses contain the same number of atoms. The number is 6.02×10^{23}.

Learning Outcomes

After going through this unit, you are expected to:

- define isotopes of atoms and calculate the relative atomic masses of various isotopic elements. Calculate the relative formula mass and the relative molecular mass of a given compound.
- calculate percentage composition and mass of elements in a given compound.
- calculate the number of moles in a given mass and the mass from a given number of moles.
- calculate the number of particles in a given mass of a substance using Avogadro's number.
- write balanced chemical equations and find stoichiometric ratios of the reactants and products.
- calculate and prepare solutions using different unit of concentration and further dilute to lower concentrations.
- identify the limiting and the excess reagents through stoichiometric calculations.
- calculate theoretical and percentage yields of substances in a given chemical reaction.

Time Frame

Suggested allotment time: **10 weeks**

If you set an average of 3 hours per day, you should be able to complete the unit comfortably by the end of the assigned week.

Try to do all the learning activities and compare your answers with the ones provided at the end of the unit. If you do not get a particular exercise right in the first attempt, you should not get discouraged but instead, go back and attempt it again. If you still do not get it right after several attempts then you should seek help from your friend or even your tutor.

DO NOT LEAVE ANY QUESTION UN-ANSWERED.

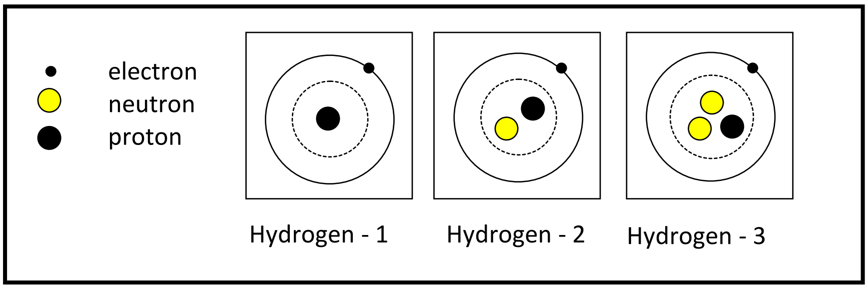
Terminologies

Before you get into the thick of things, let us make sure you know some of the terminologies that are used throughout this module.

Actual yields	Are the amount of the products that are actually produced in a reaction.
Atomic number	Is the number of protons in the nucleus of an atom.
Avogadro's Law	States that equal volumes of all gases, under the same conditions of temperature and pressure contain the same number of molecules.
Dilution	Is the process of weakening the concentration of a solute in solution by simply mixing with more solvent.
Empirical Formula	Is the simplest formula. It shows the simplest whole number ratio of the elements present in a compound. The formula obtained from percentage compositions are the empirical formulas.
Isotopes	Are atoms of the same elements with the same number of protons (atomic number) but different number of neutron or different mass number.
Mass number	Is the total number of protons and neutrons found in the nucleus of an atom.
Molar mass (M)	The mass of one mole of particles of any substance. The unit used is **grams per mole (g/mol)**.
Mole	Is the unit of measurement for atoms and molecules.
Molecular Formula	Is the true formula of a compound. It shows all the atoms present in a molecule. Most empirical formulae are also the molecular formulae.
Percent Abundance of isotopes	Is the relative percentage of a particular isotope.
Relative atomic mass (A_r)	Is the number of times the mass of one atom of an element is greater than $^{1}/_{12}$ of the mass of one carbon-12 atom.
Standard Solution	Is a solution containing known concentration of an element or a substance and a known weight of solute that is dissolved to make a specific volume. It is prepared using a standard solution called **primary standard**.
Stoichiometry (stoi-kio-me-tree)	Is the calculation of relative quantities of reactants and products in chemical reactions.
Theoretical yields	Are the amount of products calculated from the complete reaction of the limiting reactant.

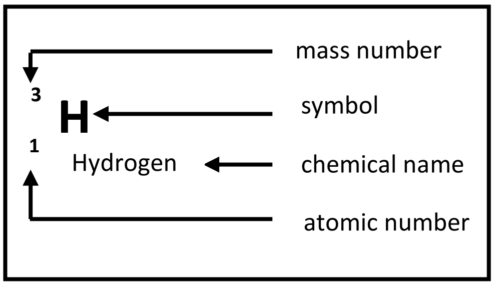
12.1.1 Isotopes

Look at the models of hydrogen atoms shown below. What are the similarities and differences about them?



They are similar except that hydrogen–1 has no neutrons, hydrogen-2 has one neutron, and hydrogen-3 has two neutrons. These hydrogen atoms are known as **isotopes**. **Isotopes** are atoms of the same elements with the same number of protons but different number of neutrons.

We can also write the symbol of one of the three isotopes of hydrogen shown below:



The nuclear symbol of the isotope hydrogen-3 or tritium.

The number of protons in the nucleus of an atom is called the **atomic number** (proton number and is given the symbol **Z**). In the symbol of hydrogen found above, hydrogen-3 has an atomic number of 1. Neutrons and protons have a similar mass.

Electrons possess very little mass. So the mass of any atom depends on the number of neutrons and protons in its nucleus. The total number of protons and neutrons found in the nucleus of an atom is called the **mass number (nucleon number).** It is given the symbol **A**. Hydrogen-3 has 2 neutrons and 1 proton in its nucleus and has a mass number of 3.

To find the mass number of hydrogen-3, we will use the relationship below:

Mass number (A)	=	atomic number (Z)	+	number of neutron
3	=	1	+	2

To find the number of neutrons of hydrogen-3, use the relationship below:

Number of neutrons = mass number (A) - atomic number
= 3 - 1
= 2

Mass number (A)	=	atomic number (Z)	+	number of neutrons

Number of neutrons	=	mass number (A)	–	atomic number (Z)

For example, the number of neutron in one atom of $^{24}_{12}Mg$

$$\underset{(A)}{24} - \underset{(Z)}{12} = 12$$

Isotopes are atoms of the same elements with the same number of protons (atomic number) but different number of neutron or different mass number.

Most elements that commonly occur are made up of isotopes. For example, chlorine consists of two isotopes. A sample of chlorine gas consists of 75% chlorine-35 and 25% chlorine-37. A few elements do not have isotopes. For example, all atoms of fluorine contain ten neutrons and nine protons.

Most isotopes have the same chemical properties, but slightly different physical properties. The chemical properties of isotopes are similar because chemical reactions involve only the electrons and not the protons and neutrons. The physical properties differ because the relative masses of the isotopes differ. For example, hydrogen-2 has a slightly higher boiling point and density than hydrogen-1.

Physical properties include mass, density, boiling point and melting point while chemical properties determine how an atom behaves during a chemical reaction.

Uses of isotopes

Isotopes that emit high energy radiation are called **radio-isotopes**. They are classified as radioactive substances. Radiation emitted by radioisotopes is dangerous because it can damage living cells and cause cancer. Some radioisotopes can have important applications and can be safely used if they are handled properly.

The table below shows some examples of elements with two and three isotopes:

Element	Isotopes		
chlorine	chlorine-35	chlorine-37	
silver	silver-107	silver-109	
antimony	antimony-123	antimony-122	
copper	copper-63	copper-65	
carbon	carbon-12	carbon-14	
lithium	lithium-6	lithium-7	
gallium	gallium-69	gallium-71	
uranium	uranium-238	uranium-235	
magnesium	magnesium-24	magnesium-25	magnesium-26
strontium	strontium-86	strontium-88	strontium-90
cobalt	cobalt-56	cobalt-58	cobalt-60
silicon	silicon-28	silicon-29	silicon-30

Examples of some elements and their isotopes

Relative atomic masses of isotopes

Atoms are very small particles. Atoms also have very small masses, so it is not practical to use the actual masses of atoms in calculations. To overcome this problem, chemists often compare masses of different atoms with carbon-12 atoms (an isotope of carbon).

Scientists all over the world agreed to give the carbon-12 atom a relative atomic mass of 12. The masses of all other atoms are compared with one-twelfth the mass of one carbon-12 atom.

The **relative atomic mass (A_r)** of any atom is the number of times the mass of one atom of an element is greater $\frac{1}{12}$ than of the mass of one carbon-12 atom.

$$\text{Relative atomic mass}(A_r) = \frac{\text{Mass of one atom of the element}}{\text{Mass of } \frac{1}{12} \text{ of an atom of carbon-12}}$$

For example, one atom of oxygen is 16 times heavier than $^1/_{12}$ of an atom of carbon-12. Oxygen has a relative mass of 16. The symbol for relative atomic mass is $\mathbf{A_r}$. Relative atomic mass is a ratio and has no unit. The relative atomic masses of elements are given in the Periodic Table (see appendix).

The table below shows the relative atomic masses (A_r) of some common elements:

Element	Relative atomic mass (A_r)
hydrogen	1
carbon	12
oxygen	16
chlorine	35.5

Relative atomic masses (A_r) of some common elements

Why some A_r values not whole numbers?

The relative atomic mass of an element is usually a whole number. For some elements, such as chlorine, are not whole numbers. This is because such elements occur as mixtures of isotopes.

Use the relationship below to calculate for the relative atomic mass of isotopes of an element in the Periodic Table:

$$\left(\frac{\textbf{Percentage abundance of isotope 1}}{\textbf{100}} \ \mathbf{x} \ \textbf{Mass number of isotope 1}\right) + \left(\frac{\textbf{Percentage abundance of isotope 2}}{\textbf{100}} \ \mathbf{x} \ \textbf{Mass number of isotope 2}\right)$$

For example, chlorine exits in two isotopic forms: chlorine-35 and chlorine-37. A sample of chlorine is made of 75% of chlorine-35 atoms and 25% of chlorine-37 atoms.

To find the relative atomic mass of chlorine is:

$$= \frac{75 \text{ x } 35}{100} + \frac{25 \text{ x } 37}{100}$$

$$= 26.25 + 9.25$$

$$= \mathbf{35.5}$$

Another example is the calculation to find the relative atomic mass of copper. Copper (Cu) has two isotopes, Copper-63 (69%) and copper-65 (30%).

$$= \frac{69 \text{ x } 63}{100} + \frac{30 \text{ x } 65}{100}$$

$$= 43.47 + 19.50$$

$$= \mathbf{62.97}$$

The following exercises show how to calculate the relative atomic mass (A_r) of some common elements:

1. A sample of gallium (Ga) contains 60% of atoms of ^{69}Ga and 40% of atoms of ^{71}Ga.

   Find the relative atomic mass of gallium.

   Answer:

$$= \frac{60 \times 69}{100} + \frac{40 \times 71}{100}$$
$$= 41.40 + 28.40$$
$$= \mathbf{69.80}$$

2. Given that the percentage abundance of ^{20}Neon (Ne) is 90% and that of ^{22}Ne is 10 %.

   Find the relative atomic mass (A_r) of neon.

   Answer:

$$= \frac{90 \times 20}{100} + \frac{10 \times 22}{100}$$
$$= 18.00 + 2.20$$
$$= \mathbf{20.20}$$

3. Antimony (Sb) consists of two isotopes. The heavier isotope which makes up about 43% of naturally occurring antimony has a mass number of 123.

   Find the percentage abundance of the lighter isotope with a mass number of a 121 and show that the relative atomic mass of antimony is 122.

   Answer:

   The only percentage given is **43 %** for the heavier isotope, so the percentage of lighter isotope would be **100 % - 43 % = 57 %.**

$$= \frac{43 \times 123}{100} + \frac{57 \times 121}{100}$$
$$= 52.89 + 68.97$$
$$= \mathbf{121.86}$$

Now, check what you have just learnt by trying out the learning activity below!

Learning Activity 1 **30 minutes**

Answer the following questions in the space provided.

For Questions 1 to 4, show your working out.

1. Carbon(C) has two isotopes: ${}^{12}_{6}C$ and ${}^{13}_{6}C$.

   How many protons, electrons, and neutrons does each isotope have?

2. Write nuclear symbols for the three isotopes of cobalt (Co) (Z = 27) in which there are 29, 31, and 33 neutrons respectively.

3. The element lithium (Li) has two isotopes, lithium-6, which has a relative isotopic mass of 6.02 and makes up 7.4% of the element and Lithium-7 which has a relative isotopic mass of 7.02 and makes 92.6% of the element.

   Calculate the relative atomic mass of lithium.

4. Use the information in the table below to calculate the relative atomic mass of silicon (Si).

Isotope	Relative isotopic mass	Relative abundance (%)
${}^{28}Si$	27.98	92.2
${}^{29}Si$	28.98	4.7
${}^{30}Si$	29.97	3.1

Isotopes of silicon (Si)

5. Write symbols to represent the following isotopes:

   a. 6 protons, 7 neutrons, and 6 electrons ___________
   b. 19 protons, 20 neutrons, and 19 electrons ___________
   c. 11 protons, 12 neutrons, and 11 neutrons ___________

6. Give the meaning of isotope.

   ____
   ____

Thank you for completing your learning activity 1. Check your work. Answers are at the end of this module.

Percentage abundance of isotopes

An isotope of an element is a variant in the nuclear make-up of the atom. Isotopes have different numbers of neutrons from atom to atom of the same element. The number of protons in the nucleus does not change from atom to atom. The mass number of every element changes from atom to atom. The **mass number** of every element is the sum of protons and neutrons found in the nucleus. Because isotopes have different number of neutrons, the mass number is not the same and is a method identification of isotopes.

The **percent abundance** of isotopes is the relative percentage of a particular isotope. If the mass numbers of the isotopes are known and the element's periodic table average mass number is known, simple algebra can be used to calculate the percent abundance.

We can use the relationship below to calculate the percentage abundance of the isotopes of any element.

$$\textbf{Average mass} = \left(\begin{array}{l}\textbf{Relative}\\ \textbf{isotopic mass}\\ \textbf{of isotope 1}\end{array} \times \begin{array}{l}\textbf{Percentage}\\ \textbf{abundance}\\ \textbf{of isotope 1}\end{array}\right) + \left(\begin{array}{l}\textbf{Relative}\\ \textbf{isotopic mass}\\ \textbf{isotope 2}\end{array} \times \begin{array}{l}\textbf{1 – Percentage}\\ \textbf{abundance of}\\ \textbf{isotope 1}\end{array}\right)$$

Example 1

Rubidium (Rb) has two isotopes. Rubidium -85 with mass number of 84.9117 and Rubidium-87 with mass number of 86.9085. The average mass is 85.4678.

Find the percentage abundance of the two isotopes of rubidium.

Solution:

Let x % for Rubidium-85 with a mass of 84.9117.

Let 1 – x for Rubidium-87 with a mass of 86.9085. In 1-x, 1 is therefore 100%.

The average atomic mass is 85.4678.
Therefore;

$$
\begin{array}{lllll}
85.4678 & = & (84.9117)\ (x) & + & (86.9085\ \times\ 1 - x) \\
85.4678 & = & 84.9117\,x & + & 86.9085\ -\ 86.9085x \\
85.4678 & - & 86.9085 & = & 84.9117x - 86.9085x \\
 & = & \dfrac{--1.4407}{-1.9968} & = & \dfrac{-1.9968x}{-1.9968} \\
 & & \mathbf{0.7215} & \mathbf{=} & \mathbf{x}
\end{array}
$$

- To get the percentage abundance of isotope 1 which is Rubudium-85, use **(x x 100%)**. So, for rubidium-85 it has **0.7215 x 100% = 72.15%**
- To get the percentage abundance of isotope 1 which is Rubidium-87, use **(100% - x).** So, **100 % - 72.15 = 27.85** .

Therefore;
Rubidium-85 has the percentage abundance of 72.15%.
Rubidium-87 has the percentage abundance of 27.85%.

Now, check what you have just learnt by trying out the learning activity below!

Learning Activity 2

40 minutes

Answer the following questions in the space provided.

Show your working out.

1. Find the percentage abundance of the two isotopes of Bromine (Br).

   Bromine-79 has a mass of 78.9183 and bromine-81 has a mass of 80.9163. The average atomic mass is 79.9041 grams per mole.

2. Lithium has 2 natural isotopes, Lithium-6 and lithium-7. They have atomic masses of 6.0151 and 7.0160 atomic mass units respectively. The average atomic mass of lithium is 6.941.

   Calculate the natural abundance of these two isotopes.

Thank you for completing your learning activity 2. Check your work. Answers are at the end of this module.

Avogadro's Number

The unit of measurement for atoms and molecules is the **mole**. The mole is the unit for chemical quantity. The symbol for the mole is **mol**. A mole of a substance contains the same number of particles as the number of atoms in 12g of carbon-12.

Amedeo Avogadro (1776 – 1856), an Italian scientist whose name is associated with the mole.

How many particles are there in a mole?
There are approximately **6.02×10^{23}** particles in one mole of substance which is called the **Avogadro's constant** or **Avogadro's number** in honour of Amedeo Avogadro.

How do we convert between number of moles and number of particles?
Since one mole of a substance is 6.02×10^{23} particles,

$$\boxed{\text{Number of moles} = \frac{\text{Number of particles}}{6.02 \times 10^{23}}}$$

Equal number of moles contains equal number of particles. The reverse is also true.

Example 1:

Convert 1 x 10^{23} of neon atoms to moles of neon atoms.
Solution:

$$\text{Number of moles} = \frac{\text{number of particles}}{6.02 \times 10^{23}}$$

$$= \frac{1 \times 10^{23}}{6.02 \times 10^{23}}$$

$$= \mathbf{0.166mol}$$

Example 2:
How many iron atoms are there in 0.5 mol of iron?

Solution:
Number of iron atoms = number of moles x 6.02×10^{23}
= 0.5 x 6.02×10^{23}
= **3.01 x 10^{23}**

Example 3:
How many hydrogen atoms are there in three moles of hydrogen gas?

Solution:
Hydrogen gas is made up of hydrogen molecule (H_2). In one mole of hydrogen molecules (H_2), there are two moles of hydrogen (H) atoms.

In three moles of hydrogen molecules (H_2), there are six moles of hydrogen (H) atoms.
= 6 x 6.02 x 10^{23}
= **3.612 x 10^{24}**

Calculating the number of particles
Use the general relationship to calculate the number of particles:

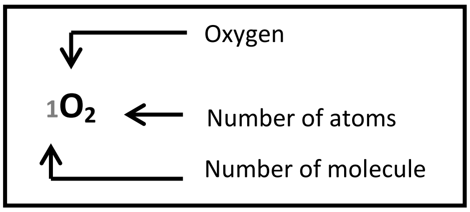
Number of particles in a given amount of substance	**=**	**Amount of substance (mol)**	**x**	**Number of particles in 1 mol (6.02 x 10^{23})**

Example 1:
Calculate the number of oxygen molecules(O_2) in 2.5mol of oxygen(O_2). By definition, 1 mole of oxygen molecules(O_2) contains 6.02×10^{23} molecules.

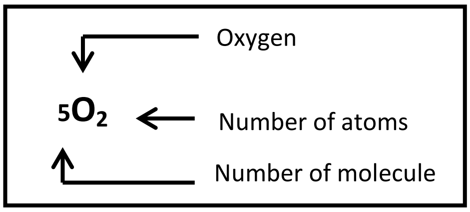
Number of oxygen molecules in 2.5mol of oxygen molecules(O_2) = 2.5 x 6.02×10^{23} = **1.505 x 10^{24}**

Example 2:

Calculate the amount in mole of oxygen atom (O) in 5 moles of oxygen molecules (O_2). One oxygen molecule (O_2) contains two atoms. 1 mole of oxygen molecules (O_2) contains 2 moles of oxygen atom (O). So, 5 moles of oxygen molecules (O_2) contain 10 moles of oxygen atom (O).



One oxygen molecule containing two moles of oxygen atoms



Five moles of oxygen molecule containing ten moles of oxygen atoms

From the two boxes above, you can already see that the number of molecules is multiplied by the number of atoms to get the total number of moles.

Now, check what you have just learnt by trying out the learning activity below!

Learning Activity 3 **60 minutes**

Answer the following questions:

Show your working out in the boxes provided.

1. Calculate the number of:

   a. atoms in 2.0 mol of sodium(Na) atoms.

   b. molecules in 0.10 moles of nitrogen molecules(N_2).

   c. atoms in 20.0 mol of carbon(C) atoms.

d. molecules in 4.2 mol of water(H_2O).

2. Calculate the number of:

a. atoms in 1.0×10^{-2} mol of iron(Fe).

b. molecules in 4.62×10^{-5} mol of carbon dioxide(CO_2).

c. atoms in 1.6×10^{-8} mol of silicon(Si).

3. Calculate the amount in moles (mol) of:

a. chlorine atom(Cl) in 0.4 mol of chlorine molecules(Cl_2)

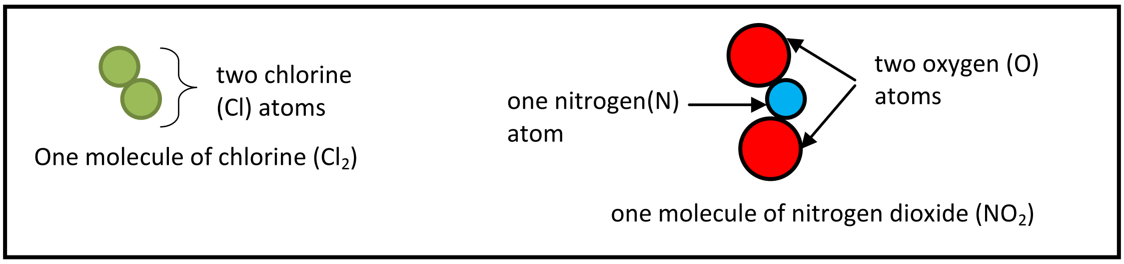
b. hydrogen atom(H) in 1.2 mol of methane(CH_4)

Thank you for completing your learning activity 3. Check your work. Answers are at the end of this module.

12.1.2 Relative Formula Mass and Percentage Composition

Relative molecular mass

Many elements and compounds exist as molecules. For example, chlorine exists as molecules. Each molecule of chlorine(Cl_2) consists of two chlorine atoms(Cl). One molecule of nitrogen dioxide(NO_2) consists of one nitrogen atom(N) and two oxygen atoms(O) as seen below:



Chlorine and nitrogen dioxide exist as a molecule

The relative molecular mass of a molecule is the average mass of the substance when compared $\frac{1}{12}$ of the mass of an atom of carbon-12.

$$\text{Relative molecular mass}(M_r) = \frac{\text{Average mass of one molecule of a substance}}{\text{Mass of } \frac{1}{12} \text{ of an atom of carbon-12}}$$

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

The **Relative Molecular Mass (RMM of M_r)** of a molecule is calculated by adding together the Relative Atomic Mass (RAM or A_r) of each atom in its chemical formula. Like relative atomic mass, it is a ratio and has no unit.

Molecule	Chemical formula	Number of atoms in a molecule	Calculating Relative Molecular Mass(M_r)
nitrogen	N_2	2N	(2 X 14) = 28
ammonia	NH_3	1N; 3H	(1 X 14) + (3 X 1) = 17
carbon dioxide	CO_2	1C; 2O	(1 X 12) + (2 x 16) = 44
water	H_2O	2H; 1O	(2 X 1) + (1 X 16) = 18
ethanol	C_2H_5OH	2C; 6H; 1O	(2 X 12) + (6 X 1) + (1 X 16) = 46

Calculating the relative molecular mass of some molecules.

Relative formula mass

Substances like water (H_2O) and carbon dioxide (CO_2) exist as molecules and are called **covalent** compounds. Substances like sodium chloride (NaCl) do not exist as molecules and are called **ionic** compounds. The relative molecular mass of an ionic compound is more accurately known as **relative formula mass**. For example, the relative formula mass of sodium chloride (NaCl) is 23 + 35.5 = 58.5.

Substance	Formula unit	Number of atoms in formula unit	Calculating Relative Formula Mass(M_r)
Magnesium sulphate	$MgSO_4$	1Mg; 1S; 4O	(1 X 24) + (1 X 32) + (4 X 16) = 120
Calcium carbonate	$CaCO_3$	1Ca; 1C; 3O	(1 X 40) + (1 X 12) + (3 X 16) = 100
Calcium nitrate	$Ca(NO_3)_2$	1Ca; 2N; 6O	(1 X 40) + (2 x 14) + (6 X 16) = 164
Copper (II) sulphate	$CuSO_4$	1Cu; 1S; 4O	(1 X 64) + (1 X 32)+ (4X16) = 160

Calculating the relative molecular mass of some ionic substances.

Calculating Relative Molecular Mass (RMM)

Example 1:
Calculate the relative molecular mass of ammonia (NH_3).

Solution:
Add the relative atomic masses of one nitrogen atom (N) and three hydrogen atoms.

RMM of ammonia (NH_3) = (1 x 14) + (3 x 1)
= 14 + 3
= **17**

Example 2:
What is the relative molecular mass of glucose ($C_6H_{12}O_6$)?

Solution:
RMM of glucose ($C_6H_{12}O_6$) = (6 x 12) + (12 x 1) + (6 x 16)
= 72 + 12 + 96
= **180**

Calculating relative formula mass

Example 1:
What is the relative formula mass of magnesium chloride ($MgCl_2$)?

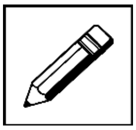
Solution:
RMM of magnesium chloride = (1 x 24) + (2 x 35.5)
= 24 + 71
= **95**

Example 2:
What is the relative formula mass of sodium hydroxide (NaOH)?

Solution:

RMM of sodium hydroxide	=	(1 x 23) + (1 x 16) + (1 x 1)
	=	23 + 16 + 1
	=	**40**

Now, check what you have just learnt by trying out the learning activity below!



Learning Activity 4

60 minutes

Answer the following questions:

For Question 1, write your answer on the spaces provided.

1. What is meant by relative molecular mass of a substance?
____
____
____

For Question 2 and 3, show your working out in the space provided

2. Calculate the relative molecular mass of each of the following molecules.

   a. hydrogen chloride (HCl)

   b. methane (CH_4)

   c. sucrose ($C_{12}H_{22}O_{11}$)

d. sulphur dioxide (SO_2)

e. sulphuric acid (H_2SO_4)

3. Calculate the relative formula mass of each of the following ionic substances.

a. calcium hydroxide ($Ca(OH)_2$)

b. potassium fluoride (KF)

c. copper(II) oxide (CuO)

d. potassium sulphide (K_2S)

e. magnesium oxide (MgO)

Thank you for completing your learning activity 4. Check your work. Answers are at the end of this module.

Molar masses

To overcome the difficulty of dealing with ionic compounds that do not form molecules, we sometimes use the term molar mass.

The mass of one mole of particles of any substance is called the **molar mass** (M). The unit used is **grams per mole (g/mol)**.

Do you notice a relationship between the value of relative atomic mass (A_r) and the molar mass (M) of a substance? The molar mass is equal to the relative atomic mass (RAM) of the element in grams (g).

For example, the relative atomic mass of sodium (Na) is 23. The mass of one mole of sodium atoms is 23g. We can also say that the molar mass of sodium is 23grams per mole (23 g/mol).

Element	Relative atomic mass (A_r)	Molar mass
aluminium	27	27g/mol
carbon	12	12g/mol
neon	20	20g/mol
oxygen	16	16g/mol

The molar masses of some elements.

What is the relationship between mole and molar masses?
The number of moles of an element can be calculated using the formula:

$$\textbf{Number of moles of element (n)} = \frac{\textbf{Mass of element in grams (g)}}{\textbf{Relative atomic mass }(\mathbf{A_r})\textbf{ of element}}$$

Example 1:
Determine the number of moles of 4.4g of carbon dioxide (CO_2).

Solution:

$$\text{Number of moles of element (n)} = \frac{\text{Mass of element in grams (g)}}{\text{Relative atomic mass }(A_r)\text{ of element}}$$

$$= \frac{4.4}{44}$$

$$= \mathbf{0.1\ mol}$$

Example 2:
How many moles of lead are there in 20.7g of lead?

Solution:

$$\text{Number of moles of element (n)} = \frac{\text{Mass of element in grams (g)}}{\text{Relative atomic mass }(A_r)\text{ of element}} = \frac{20.7}{207}$$

$$= \mathbf{0.1\ mol}$$

What is the mass of one mole of molecule or one mole of a compound?

You have learnt that the mass of one mole of atoms is the same as the relative atomic mass (A_r) in grams (g). The same idea can be extended to molecule and compound. One mole of a substance will have a mass equal to the relative molecular mass or relative formula mass in grams.

Substance	Formula	Relative molecular mass	Molar mass
oxygen	O_2	(2 x 16) = 32	32g/mol
iodine	I_2	(2 x 127) = 254	254g/mol
magnesium fluoride	MgF_2	(1 x 24) + (2 x 19) = 62	62g/mol
water	H_2O	(2 x 1) + (1 x 16) = 18	18g/mol

Calculating the molar mass of some common substances.

The number of moles of substances can be calculated using this formula:

$$\text{Number of moles} = \frac{\text{Mass of a substance (g)}}{\text{Relative Formula Mass (RFM)}}$$

Example1:

Find the number of moles of 88g carbon dioxide (CO_2), using the formula above.

$$\text{Number of moles of carbon dioxide} = \frac{88}{44}$$

$$= \textbf{2 moles of carbon dioxide}$$

Example 2:

Find the number of moles of 175.5g sodium chloride (NaCl), using the formula above.

$$\text{Number of moles of sodium chloride} = \frac{175.5}{58.5}$$

$$= \textbf{3 moles of sodium chloride}$$

Remember, we have also learnt that:

$$\textbf{Number of moles} = \frac{\textbf{Number of particles}}{\mathbf{6.023 \times 10^{23}}}$$

Example 3:
Convert 2.01 x 10 ^{23} magnesium atom to mole of magnesium atoms.

$$\text{Number of moles carbon dioxide} = \frac{2.01 \times 10^{23}}{6.023 \times 10^{23}}$$

$$= \mathbf{0.334 \text{ mole of magnesium atom}}$$

Now, check what you have just learnt by trying out the learning activity below!

Learning Activity 5

60 minutes

Answer the following questions:

Show your working out.

1. What is the mass of one mole of the following molecules?

   a. hydrogen(H_2)

   b. oxygen(O_2)

   c. nitrogen(N_2)

   d. chlorine(Cl_2)

2. Calculate the number of moles of molecules in the following:

   a. 32g of oxygen(O_2)

   b. 7g of nitrogen(N_2)

c. 8g of hydrogen(H_2)

d. 71g of chlorine(Cl_2)

3. Calculate the mass of each of the following:

   a. 7mol of nitrogen(N_2)

   b. 3mol of water(H_2O)

   c. 4 mol of hydrogen chloride(HCl)

   d. 0.5mol of sulphur dioxide(SO_2)

Thank you for completing your learning activity 5. Check your work. Answers are at the end of this module.

Percentage by mass or percentage composition

When buying certain products such as food stuff and medicines, we often look at the labels for their ingredients and in what proportions they are present.

The **percentage composition** of a compound tells us the percentage mass of each element in the formula.

How do we find the percentage composition of a substance?
In general, the percentage by mass of an element in a compound can be found using this formula:

Example 1

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

Find the percentage by mass of hydrogen atom(H) and oxygen atom(O) in hydrogen peroxide(H_2O_2).

Solution:
(i) The relative molecular mass of hydrogen peroxide (H_2O_2) is (2 x 1) + (2 x 16) = **34.**
(ii) The percentage of hydrogen atoms (H) in hydrogen peroxide (H_2O_2) is:

$$\%\text{ by mass of hydrogen (H) in hydrogen peroxide } (H_2O_2) = \frac{A_r \text{ of hydrogen (H) } \times \text{ number hydrogen of atoms in formula}}{M_r \text{ of hydrogen peroxide } (H_2O_2)} \times 100\%$$

$$= \frac{\mathbf{1 \times 2}}{\mathbf{34}} \times \mathbf{100\%}$$

$$= \mathbf{5.9\%}$$

(iii) The percentage of oxygen atoms (O) in hydrogen peroxide (H_2O_2) is:

$$\%\text{ by mass of oxygen (O) in hydrogen peroxide } (H_2O_2) = \frac{A_r \text{ of oxygen (H) } \times \text{ number oxygen of atoms in formula}}{M_r \text{ of hydrogen peroxide } (H_2O_2)} \times 100\%$$

$$= \frac{\mathbf{16 \times 2}}{\mathbf{34}} \times \mathbf{100\%}$$

$$= \mathbf{94.1\ \%}$$

Check total percentage $= \mathbf{5.9\ \% + 94.1\ \%}$

$$= \mathbf{100\%}$$

Example 2:
Calculate the percentage composition of each element in calcium sulphate ($CaSO_4$).

Solution:
(i) Relative formula mass of calcium (II) sulphate ($CaSO_4$):

Calcium (Ca)	= (1 x 40)	=	40
Sulphur (S)	= (1 x 32)	=	32
Oxygen (O)	= (4 x 16)	=	64
RMM of calcium (II) sulphate ($CaSO_4$)		=	**136**

(ii) % of calcium(Ca) $= \frac{40}{136} \times 100\%$

$= \mathbf{29.4\%}$

% of sulphur(S) $= \frac{32}{136} \times 100\%$

$= \mathbf{23.5\%}$

% of oxygen(O) $= \frac{64}{136} \times 100\%$

$= \mathbf{47.1\%}$

Check total percentage = **29.4 % + 23.5 % + 47.1 %**

= **100 %**

Now, check what you have just learnt by trying out the learning activity below!

Learning Activity 6

60 minutes

Answer the following questions:

Show your working out in the boxes provided.

1. Iron(Fe) can be obtained from an iron ore called haematite or iron(III) oxide which has the formula Fe_2O_3.

   Calculate the percentage of iron in haematite (Fe_2O_3).

2. Potassium chloride(KCl) and potassium nitrate(KNO_3) can be used as fertilizers to increase the potassium content of soil.

   Which salt will contain a higher percentage of potassium by mass?

3. Calculate the percentage of water in hydrated copper(II) sulphate($CuSO_4.5H_2O$).

4. Calculate the following:

   a. The percentage of nitrogen(N) in potassium nitrate(KNO_3).

   b. The percentage of chlorine(Cl) in ammonium chloride(NH_4Cl).

Thank you for completing your learning activity 6. Check your work. Answers are at the end of this module.

12.1.3 Moles

As atoms are very small, we deal with a very large number of atoms in chemistry. The mole (mol), given the symbol **n** is the **SI (Systemae Internationale)** unit for the amount of substance. We discussed earlier that the actual number of particles in a mole has been experimentally determined to be 6.02×10^{23} particles representing atoms, molecules, and ions or formula units.

For example, a mole of copper (Cu) has 6.02×10^{23} atoms, a mole of water (H_2O) has 6.02×10^{23} molecules, and a mole of sodium chloride (NaCl) has 6.02×10^{23} formula units.

Amount in Moles of a Mass of Substance

As we have learnt in our previous lesson, mole and molar masses have a relationship to each other since the relative atomic mass is equal to one mole of a particular element and the total relative molecular masses of each atom in a compound or molecule is equal to its molar mass measures in grams per mole (g/mol).

Since **molar mass** is defined as the mass (in grams) of one mole (n) of a substance, we will use again the formula below:

$$\boxed{\text{Number of moles}(n) = \frac{\text{Mass of a substance}(g)}{\text{Relative atomic mass}(A_r)\text{ or relative molecular mass}(M_r)}}$$

To simplify the formula, we will use the one below:

$$\boxed{\mathbf{n = \frac{m}{A_r \text{ or } M_r}}}$$

Where **n** is called the number of moles; **m** is the mass given; $\mathbf{A_r}$ is the relative atomic mass and $\mathbf{M_r}$ the relative molecular mass.

Example 1:
How many moles are there in 112g of iron (Fe)?

Solution: $n = \frac{m}{A_r}$

$= \frac{112}{56}$

$=$ **2mol**

Example 2:
What is the amount in moles of sodium (Na) present in 4.6g of sodium?

Solution: $n = \frac{m}{A_r}$

$= \frac{4.6}{23}$

$=$ **0.2mol**

Example 3:
What is the amount in moles of aluminium (Al) present in 9.0g of aluminium?

Solution: $n = \frac{m}{A_r}$

$= \frac{9.0}{27}$

$=$ **0.33mol**

Example 4:
What is the amount in moles of 20g sodium hydroxide?

Solution: $n = \frac{m}{A_r}$

$= \frac{20}{40}$

$=$ **0.5mol**

Now, check what you have just learnt by trying out the learning activity below!

Learning Activity 7 **60 minutes**

Answer the following questions:

Show your working out in the boxes provided.

1. Calculate the number of moles of each of the following elements:
   a. 100g of calcium (Ca)

b. 3.9g of potassium (K)

c. 70g of silicon (Si)

d. 9.5g of fluorine (F)

2. Find the amount in moles of the following compounds:

a. 19.5g of sodium chloride (NaCl)

b. 267g of aluminium chloride($AlCl_3$)

c. 480g of copper (II) sulphate($CuSO_4$)

d. 42.5g of silver nitrate($AgNO_3$)

e. 42.4g of sodium carbonate(Na_2CO_3)

Thank you for completing your learning activity 7. Check your work. Answers are at the end of this module.

Amount in moles of solutions

Chemists often need to know the concentration of a solution. Sometimes it is measured in grams per cubic decimetre (g/dm^3) but more often concentration is measured in moles per cubic decimetre (mol/dm^3).

When 1 mole of a substance is dissolved in water and the solution is made up to 1 cubic decimetre ($1dm^3 = 1000cm^3$) or 1 liter (L = 1000mL) then a **1 molar (1M) solution** is produced.

Chemists do not always need to make up such large volumes of solution. A simple method of calculating the concentration is by using the relationship:

$$\boxed{\textbf{Concentration } (mol/dm^3) = \frac{\textbf{Number of moles (n)}}{\textbf{Volume } (dm^3)}}$$

Simply, we can use;

$$\mathbf{c = \frac{n}{v}}$$

Where, **C** is the concentration of solution, **n** is the number of moles, and **V** the volume measured in cubic decimetre(dm^3) or in litres(L).

Sometimes chemists need to know the mass of a substance that can be dissolved to prepare a known volume of solution at a given concentration. A simple method of calculating the number of moles in a solution is by using the relationship below:

$$\boxed{\text{Number of moles} = \frac{\text{Concentraton}(mol/dm^3)}{\text{Volume of solution}(dm^3)}}$$

Simply, we can use; $\boxed{\mathbf{n = C \times V}}$

Example 1:
Calculate the concentration (in mol/dm^3) of a solution of sodium hydroxide (NaOH), which is made by dissolving 10g of solid sodium hydroxide (NaOH) in water and making up to 0.25 dm^3.

Solution:

(i) 1 mole of sodium hydroxide (NaOH) contains 1 mole of sodium (Na), 1 mole of oxygen (O) and 1 mole of hydrogen (H). The molar mass is; (1 x 23) + (1 x 16) + (1 x 1) = **40g**.

(ii) Find the number of moles of 40g sodium hydroxide.

$$= \frac{10}{40}$$

$$= \mathbf{0.25mol}$$

(iii) Find the concentration of sodium hydroxide solution using:

$$C = \frac{n}{V}.$$

$$= \frac{0.25}{0.25}$$

$$= \textbf{1 mol/dm}^{\mathbf{3}} \textbf{ or 1M (M stands for molar solution)}$$

Example 2:
Calculate the mass of potassium hydroxide (KOH) that needs to prepare 0.5dm^3 of a 2 mol/dm^3 (2M) solution in water.

Solution:

(i) Find the number of moles of potassium hydroxide (KOH) using the formula n = C x V.

n = C x V
= 2 x 0.5
= **1 mol**

(ii) Find the mass of potassium hydroxide (KOH).

m = n x M_r (mass of 1 mole of potassium hydroxide (KOH)
= 1 x 56
= **56g**

Note: The mass of 1 mole of potassium hydroxide (KOH)
= (1 x 19) + (1 x 16) + (1 x 1)
= **56**

Now, check what you have just learnt by trying out the learning activity below!

Learning Activity 8 **60 minutes**

Answer the following questions:

Show your working out in the space provided.

1. Calculate the concentration of the following solutions:

a. 0.50dm^3 of solution which contains 0.24mol of glucose solution ($C_6H_{12}O_6$)

b. $0.20dm^3$ of solution which contains 0.010mol of sodium chloride (NaCl)

c. $25.0dm^3$ of solution which contains 2×10^{-3}mol of solute.

d. $4.1dm^3$ of solution which contains 1.23mol of solute.

2. Calculate the number of moles present in the following solutions:

a. 1.0L of 3.4M ammonia (NH_3) solution.

b. 0.45L of 0.24M ethanol (C_2H_5OH) solution.

c. $0.015dm^3$ of 6.0M ethanol (C_2H_5OH) solution.

d. $0.015dm^3$ of 2.5M sucrose ($C_{12}H_{22}O_{11}$) solution.

Thank you for completing your learning activity 8. Check your work. Answers are at the end of this module.

Amount in moles of a sample of gas at standard temperature and pressure (s.t.p.) and at room temperature and pressure (r.t.p.)

You have learnt that the mass of one mole of a substance has a numerical value equal to its relative atomic mass or relative molecular mass. This property is applicable to solids, liquids 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 since gases weigh very little. Is there a way to relate moles to the volumes of gases?

Avogadro's Law states that equal volumes of all gases, under the same conditions of temperature and pressure contain the same number of molecules.

Chemists have found that **one mole of any gas occupies 24dm^3** (24,000cm^3) at r.t.p or room temperature (25^{o}C) and pressure (1 atmosphere). This volume is called the **molar volume** of a gas.

This means that at r.t.p.
- 1mol of oxygen occupies 24dm^3.
- 1mol of carbon dioxide occupies 24dm^3.
- 2mol of oxygen occupy 2 x 24 = 48dm^3.
- 2mol of carbon dioxide occupy 2 x 24 = 48dm^3.

NOTE: **As you go to this topic, you will be using r.t.p. for room temperature and pressure and s.t.p. for standard temperature and pressure.**

How can we calculate the number of moles of a gas?
The number of moles of a gas can be measured in three ways:

1. Find the mass of a gas. Then use the following formula to calculate the number of moles of a gas.

$$\boxed{\textbf{Number of moles gas} = \frac{\textbf{Mass of gas in grams}}{\textbf{M}_r\textbf{ of gas}}}$$

Simply, we can use;

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

2. Find the volume of gas at room temperature and pressure (r.t.p.) Then use this formula.

$$\text{Number of moles gas} = \frac{\text{Volume of gas in cm}^3 \text{ at r.t.p. } (25^0\text{C})}{24{,}000\text{cm}^3}$$

Simply, we can use;

$$\boxed{\mathbf{n = \frac{V \text{ at r.t.p.}}{24{,}000cm^3}}}$$

3. Find the volume of gas at standard temperature and pressure (s.t.p.). Use the formula below.

$$\textbf{Number of moles gas} = \frac{\textbf{Volume of gas in cm}^3 \textbf{ at s.t.p. } (0^0\textbf{C})}{\textbf{22,400cm}^3}$$

Simply, we can use;

$$n = \frac{V \text{ at s.t.p.}}{22{,}400\text{cm}^3}$$

This formula can be rearranged to give,

Volume of gas (in cm^3) = number of moles x 24 000
Volume of gas (in dm^3) = number of moles x 24

or

$$V = n \times 24\,000$$
$$V = n \times 24$$

Example 1:
What is the volume, in dm^3, of 8g oxygen gas (O_2) at r.t.p?

Solution:

(i) RMM of oxygen (O_2) = 2 x 16
= **32**

(ii) number of moles of oxygen $= \frac{m}{M_r}$

$= \frac{8}{32}$

= **0.25mol**

(iii) volume of oxygen = n x 24
= 0.25 x 24
= **6dm³**

Example 2:
In an experiment, hydrochloric acid reacted with calcium carbonate at room temperature and pressure of which 80cm^3 of carbon dioxide was produced.

Find the number of moles of carbon dioxide given off.

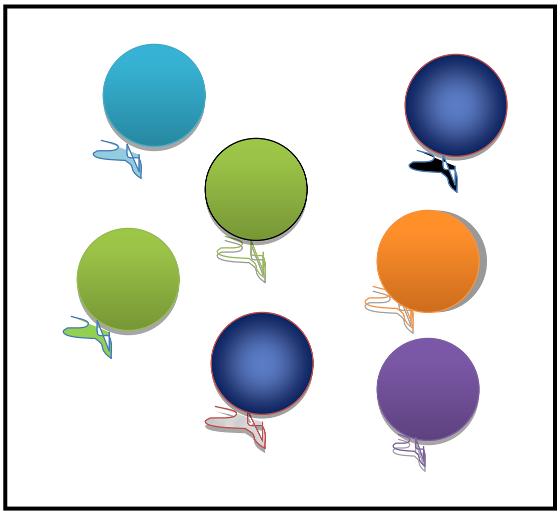
Solution:

Number of moles of carbon dioxide(CO_2) $= \dfrac{80}{24000}$

$= \mathbf{3.33 \times 10^{-3} mol}$

Look at the balloons on the right. Do balloons of the same volume contain the same number of particles?

Yes, they do! According to Avogadro's Law, each of these balloons contains the same number of gaseous particles since they have the same volume.



Balloons containing identical volumes of gas.

Do the balloons of the same mass contain the same number of particles?

If the balloons each contained 0.18 g of a different gas instead of having the same volume, they would not contain the same number of gaseous particles. Look at the table below to explain why.

Calculating the number of moles of different gases with equal masses

Gas	Mass (g)	A_r or M_r	Number of moles = $\frac{\text{mass (g)}}{A_r \text{ or } M_r}$
Helium (He)	0.18	4	$\frac{0.18}{4}$ = 0.045
Hydrogen (H_2)	0.18	2	$\frac{0.18}{2}$ = 0.090
Methane (CH_4)	0.18	16	$\frac{0.18}{16}$ = 0.011

Using the table above, you can now use the same mass of different gases to calculate the number of particles as shown in the examples below:

- Number of helium (He) atoms = $0.045 \times 6.02 \times 10^{23}$ = $\mathbf{2.70 \times 10^{22}}$
- Number of hydrogen molecules (H_2) = $0.090 \times 6.02 \times 10^{23}$ = $\mathbf{5.40 \times 10^{22}}$

Therefore:

Equal masses of different gases do not contain the same number of particles.

Now, can you calculate the number of particles in methane (CH_4) molecules using the same mass of 0.18g?

If your answer is $\mathbf{6.622 \times 10^{21}}$, then your answer is correct.

You have learnt in the previous lesson that, at s.t.p. or standard temperature(0^oC) and pressure(1 atmosphere), one mole of gas occupies a volume of $22.4dm^3$ or $22,400cm^3$ which also means that one mole of gas at s.t.p. contains 6.02×10^{23} particles.

The number of moles of gas present in a given volume at s.t.p. may be found from this formula:

$$\text{Number of moles(n)} = \frac{\text{Volume of gas in } dm^3 \text{ at s.t.p.}}{22.4dm^3}$$

$$\textbf{Number of moles of gas} = \frac{\textbf{Volume of gas in } \mathbf{dm^3} \textbf{ at s.t.p.}}{\mathbf{22.4\ dm^3}}$$

Simply, we can use: $n = \frac{\text{V at s.t.p.}}{22.4dm^3}$

The formula on the previous page can be rearranged to give the following:

Volume of gas (in cm^3)	**=**	**number of moles x 22 400**
Volume of gas (in dm^3)	**=**	**number of moles x 22.4**

or

$$\mathbf{V = n \times 22\,400}$$
$$\mathbf{V = n \times 22.4}$$

Example 1:
How many moles of hydrogen gas (H_2) are present in $11.2dm^3$ at s.t.p? How many grams does this represent?

Solution:

(i) Find the moles: $n = \frac{\text{V at s.t.p.}}{22.4}$

$$= \frac{11.2}{22.4}$$

$$= 0.5\text{mol}$$

(ii) Find the mass: $m = n \times M_r$

$$= 0.5 \times 2$$

$$= \mathbf{1g}$$

Example 2:
Calculate the volume at s.t.p occupied by 6.8g of ammonia (NH_3) gas.

(i) Find the number of moles of ammonia (NH_3) since the mass of ammonia is given.

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

$$n = \frac{6.8}{17}$$

$$= \mathbf{0.4mol}$$

Note: The molar mass of ammonia (NH_3) is (1 x 14) + (3 x 1) = 17g.

(ii) Find the volume of ammonia (NH_3) at s.t.p.

$$V = n \times 22.4$$
$$= 0.4 \times 22.4$$
$$= \mathbf{8.96dm^3}$$

Now, check what you have just learnt by trying out the learning activity below!

Learning Activity 9 **60 minutes**

Answer the following questions:

Show your working out in the boxes provided.

1. How many moles are present in the following volumes of gases at r.t.p.?

   a. $1.2dm^3$ of sulphur dioxide(SO_2)

   b. $0.24dm^3$ of methane(CH_4)

   c. $120cm^3$ of carbon dioxide(CO_2)

   d. $0.5dm^3$ of sulphur dioxide(SO_2)

2. What volume in dm^3 do the following gases occupy at r.t.p.?

   a. 0.1mol of oxygen gas(O_2)

   b. 3mol of hydrogen(H_2)

   c. 5mol of chlorine(Cl_2)

d. 0.05mol of sulphur dioxide(SO_2)

3. What volume in cm^3 do the following gases occupy at s.t.p.?

a. 0.5mol of oxygen gas(O_2)

b. 2mol of hydrogen(H_2)

c. 3mol of chlorine(Cl_2)

d. 0.25mol of sulphur dioxide(SO_2)

4. How many moles are present in the following volumes of gases at s.t.p.?

a. 1200cm^3 of sulphur dioxide(SO_2)

b. 124cm^3 of methane(CH_4)

c. 1.50dm^3 of carbon dioxide(CO_2)

d. 240cm^3 of sulphur dioxide(SO_2)

Thank you for completing your learning activity 9. Check your work. Answers are at the end of this module.

12.1.4 Empirical and Molecular Formulas

Empirical formulas

The **empirical formula** is the simplest formula. It shows the simplest whole number ratio of the elements present in a compound. The formula obtained from percentage compositions are the empirical formulas.

Some examples of empirical formulas are H_2O (water), CO_2 (carbon dioxide), H_2SO_4 (sulphuric acid or hydrogen sulphate), CH_2 and CH_3 (empirical formulas of hydrocarbon from percentage by mass of carbon and hydrogen).

The following are not empirical formulas: C_3H_6 (propene) and C_2H_6 (ethane) because they can be simplified to CH_2 and CH_3, respectively.

Finding empirical formula from percentage elemental composition

Example 1:
A hydrocarbon consists of 85.7% carbon (C) and 14.3% of hydrogen (H) by mass.

Find the empirical formula of the compound, given the relative atomic mass of carbon (C) = 12 and hydrogen (H) =1.

Solution:

	C	:	H
Ratio of mass	85.7	:	14.3
Divide by atomic mass (A_r)	$\frac{85.7}{12}$	:	$\frac{14.3}{1}$
Divide by the smaller number	$\frac{7.14}{7.14}$	:	$\frac{14.3}{7.14}$
Ratio of atoms	**1**	**:**	**2**

The empirical formula is $\mathbf{C_1H_2}$. Since 1 is never used in writing a formula, you should write $\mathbf{CH_2}$ for the empirical formula of the hydrocarbon.

Example 2:
Nicotine is a highly toxic substance found in tobacco. Analysis shows that nicotine contains 74.1% carbon (C), 8.66% hydrogen (H) and 17.3% nitrogen (N).

Determine the empirical formula of nicotine, given the relative atomic mass of carbon (C) = 12, hydrogen (H) = 1 and nitrogen (N) = 14.

Solution:

	C	:	H	:	N
Ratio of mass	74.1	:	8.66	:	17.3
Divide by atomic mass (A_r)	$\frac{74.1}{12}$	:	$\frac{8.66}{1}$	:	$\frac{17.3}{14}$
Divide by the smaller number	$\frac{6.18}{1.24}$	:	$\frac{8.66}{1.24}$	:	$\frac{1.24}{1.24}$
Ratio of atoms	**5**	**:**	**7**	**:**	**1**

The empirical formula of nicotine is $\mathbf{C_5H_7O}$.

Finding empirical formula from mass elemental composition

Example 1:
An analysis of 50g of water shows it has 5.6g hydrogen (H) and 44.4g of oxygen (O).

Calculate the empirical formula of water, given the relative atomic mass of hydrogen (H) = 1 and oxygen (O) = 16.

Solution:	H	:	O
Ratio of mass	5.6	:	44.4
Divide by atomic mass (A_r)	$\frac{5.6}{1}$	:	$\frac{44.4}{16}$
Divide by the smaller number	$\frac{5.6}{2.8}$	:	$\frac{2.8}{2.8}$
Ratio of atoms	**2**	**:**	**1**

The empirical formula of water is $\mathbf{H_2O}$.

Example 2:
An 8.4g sample of hydrogen (H) (containing carbon and hydrogen atoms only) contains 7.2g of carbon (C).

What is the empirical formula of this compound? (Relative atomic of hydrogen (H) = 1 and carbon (C) = 12).

Solution:
The mass of the sample is 8.4g. The mass of carbon is 7.2g and the mass of hydrogen is unknown.

Therefore, 8.4 – 7.2 = 1.2. So the mass of hydrogen is **1.2g**.

	C	:	H
Ratio of mass	7.2	:	1.2
Divide by atomic mass (A_r)	$\frac{7.2}{12}$	:	$\frac{1.2}{1}$
Divide by the smaller number	$\frac{0.6}{0.6}$	:	$\frac{1.2}{0.6}$
Ratio of atoms	**1**	**:**	**2**

The empirical formula of the hydrocarbon is $\mathbf{CH_2}$.

Now, check what you have just learnt by trying out the learning activity below!

Learning Activity 10 **60 minutes**

Answer the following questions:

Write your answer in the space provided.

1. A 23.9g sample of a compound contains 2.4g of carbon (C), 0.2g of hydrogen (H) and 21.3g of chlorine (Cl).

   What is the empirical formula of this compound? (Relative atomic mass of carbon (C) = 12, hydrogen (H) = 1 and chlorine (Cl) = 35.5).

2. A pure sample of an oxide of nitrogen (N) contains 1.58g of nitrogen and 3.62g of oxygen (O).

   Determine the empirical formula of the oxide, given the relative atomic mass of nitrogen (N) = 14 and oxygen (O) = 16.

3. Chemical analysis shows that naphthalene (sold commercially as moth balls) consists of 93.7% carbon (C) and 6.3% of hydrogen (H).

   Use this information to determine the formula of naphthalene given the relative atomic mass of carbon (C) = 12 and hydrogen (H) = 1.

4. A compound contains 46.7% of silicon (Si) and 53.3% oxygen (O) by mass.

   Find the empirical formula of the compound, given the relative atomic mass of silicon (Si) = 28 and oxygen (O) = 16.

5. What is empirical formula?
   ____
   ____
   ____

Thank you for completing your learning activity 10. Check your work. Answers are at the end of this module.

Molecular formulas

The **molecular formula** is the true formula of a compound. It shows all the atoms present in a molecule. Most empirical formulae are also the molecular formulae. Some examples of molecular formulae are H_2O (water), CO_2 (carbon dioxide), H_2SO_4 (sulphuric acid or hydrogen sulphate), and $Cu(NO_3)_2$ (copper nitrate). Most empirical formulas are **molecular formulas**.

Finding molecular formula from empirical ratio of elements
You can use the relationship below to find the molecular formula from the empirical formula:

Empirical formula = A_xB_y
Molecular formula = $(A_xB_y)_n$

where, **n** represents a number such as 1,2,3 and so on. To find n, the relationship below is used:

$$n = \frac{\text{relative molecular mass}}{\text{empirical formula mass}}$$

Example 1:
The fuel in disposable cigarette lighters is butane. The empirical formula of butane is C_2H_5 and its molar mass is 58g/mol.

Find the molecular formula of butane, given the relative atomic mass of carbon (C) =12, and hydrogen (H) = 1.

Solution:
(i) The empirical formula is $\mathbf{C_2H_5}$. Its mass is (2 x 12) + (5 x 1) = **29**.
(ii) Since the molar mass of butane is 58g/mol, divide it by the empirical formula mass of 29 .

The formula to use is: $\dfrac{\text{molar mass (M)}}{\text{Empirical formula mass}}$

$$\text{So, } \frac{58}{29} = 2$$

(iii) Therefore, the molecular formula is $\mathbf{(C_2H_5)_2 = C_4H_{10}}$.

Example 2:
Propene has the empirical formula CH_2. The relative molecular mass of propene is 42.

Find the molecular formula of propene. The relative atomic mass of carbon (C) = 12 and hydrogen (H) = 1.

Solution:

(i) The empirical formula is $\mathbf{CH_2}$. Its mass is (1 x 12) + (2 x 1) = **14**.

(ii) Since the molar mass of butane is 42g/mol, divide it by the empirical formula mass of 14.
So, $\dfrac{42}{14} = 3$

(iii) Therefore, the molecular formula is $\mathbf{(CH_2)_3 = C_3H_6}$.

Finding molecular formula from mass elemental composition

Example 1:
A hydrocarbon consists of 85.7% carbon (C) and 14.3% of hydrogen (H) by mass, given the relative atomic mass of carbon (C) = 12 and hydrogen (H) =1.

a. Find the empirical formula of the compound.
b. If the relative molecular mass of the compound is 56, find its molecular formula.

Solution:	C	:	H
Ratio of mass	85.7	:	14.3
Divide by atomic mass (A_r)	$\frac{85.7}{12}$	:	$\frac{14.3}{1}$
Divide by the smaller number	$\frac{7.14}{7.14}$	:	$\frac{14.3}{7.14}$
Ratio of atoms	**1**	**:**	**2**

a. Therefore, the empirical formula is $\mathbf{CH_2}$. The empirical formula mass is (1 x 12) + (2 x 1) = **14**.

b. Since the molar mass of the compound is 56g/mol, divide it by the empirical formula mass of 14. So, $\frac{56}{14} = 4$

Therefore, the molecular formula is $\mathbf{(CH_2)_4 = C_4H_8}$.

Example 2:

A hydrocarbon consists of 80% carbon and 20% hydrogen by mass. The relative molecular mass of the hydrocarbon is 30, given its relative atomic mass of carbon (C) =12 and hydrogen (H) = 1.

a. Find the empirical formula of the hydrocarbon.
b. Find the molecular formula of the compound.

Solution:	C	:	H
Ratio of mass	80%	:	20%
Divide by atomic mass (A_r)	$\frac{80}{12}$	:	$\frac{20}{1}$
Divide by the smaller number	$\frac{6.667}{6.667}$	:	$\frac{20}{6.667}$
Ratio of atoms	**1**	**:**	**2.9 rounded off to 3**

a. Therefore, the empirical formula is $\mathbf{CH_3}$. The empirical formula mass is (1 x 12) + (3 x 1) = **15**.

b. Since the molar mass of the compound is 30g/mol, divide it by the empirical formula mass of 15. So, $\frac{30}{15} = \mathbf{2}$.

Therefore, the molecular formula is $\mathbf{(CH_3)_2 = C_2H_6}$.

Now, check what you have just learnt by trying out the learning activity below!

Learning Activity 11 **60 minutes**

Answer the following questions:

For Questions 1 to 4, show your working out in the boxes and for Question 5, write your answer on the spaces provided.

1. An 8.4g sample of hydrogen (H) (containing carbon and hydrogen atoms only) contains 7.2g of carbon (C). (Relative atomic of hydrogen (H) = 1 and carbon (C) = 12)

   a. What is the empirical formula of this compound?

   b. If this compound has a molar mass of 84, what is its molecular formula?

2. The empirical formula of ethane (used to make polyethylene plastic) was found by experiment to be CH_2. The molar mass of this substance was determined to be 28 g/mol.

   Find the molecular formula of ethane. The relative atomic mass of carbon (C) = 12 and hydrogen (H) = 1.

3. Compound X contains 40.0% carbon, 6.6% hydrogen and 53.3% oxygen. Its relative molecular mass is 180.

   What is the molecular formula of compound X?

4. Caffeine is a compound found in coffee and tea. The percentage composition of caffeine is 49.1% carbon, 5.1% hydrogen, 16.5% oxygen and 28.9% nitrogen. The relative molecular mass of caffeine is 195.

   Find the molecular formula of caffeine, given the atomic mass of carbon (C) =112, hydrogen (H) = 1 and oxygen (O) = 16, and nitrogen (N) = 14.

5. What is molecular formula?
   ____
   ____
   ____

Thank you for completing your learning activity 11. Check your work. Answers are at the end of this module.

12.1.5 Stoichiometry

Stoichiometry (stoi-kio-me-tree) is the calculation of relative quantities of reactants and products in chemical reactions. Stoichiometry is found on the **Law of Conservation of Mass** where the total mass of the reactants equals the total mass of the products. This means that if the amounts of the separate reactants are known, then the amount of the product can be calculated. Also, if one reactant has a known quantity and the quantity of the product can be determined, then the amount of the other reactants can also be calculated.

For example, when methane (CH_4) reacts with oxygen gas (O_2), carbon dioxide (CO_2) and water are formed as in balanced equation below:

$$\text{Methane} + \text{oxygen gas} \longrightarrow \text{carbon dioxide} + \text{water}$$

$$CH_{4(g)} + 2O_{2(g)} \longrightarrow CO_{2(g)} + 2H_2O_{(l)}$$

From the equation, one molecule of methane reacts with two molecules of oxygen gas to form one molecule of carbon dioxide and two molecules of water.

Stoichiometry measures these quantitative relationships, and is used to determine the amount of products or reactants that are produced or needed in a given reaction.

Masses of reactant and product given the mass of one reactant

We have learnt how to construct balanced chemical equation. The substances which react are called the **reactants** and the substances which are formed are called the **products**. We have also learnt that a chemical equation represents the actual atoms or molecules taking part in a reaction.

From a balanced chemical equation, we are able to calculate the amount of reactants required and the amount of products formed as well.

Let us consider the reaction between zinc (Zn) and hydrochloric acid (HCl). The products formed are zinc chloride ($ZnCl_2$) and hydrogen gas (H_2).

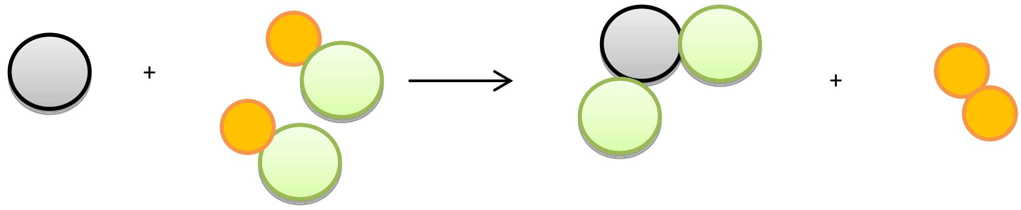
The chemical equation for the reaction is:

$$\text{zinc} + \text{hydrochloric acid} \longrightarrow \text{zinc chloride} + \text{hydrogen gas}$$

$$Zn_{(s)} + 2HCl_{(aq)} \longrightarrow ZnCl_{2(aq)} + H_{2(g)}$$

The graphical illustration of the reaction between zinc and hydrochloric acid is shown in the next page.

Graphical illustration:



$Zn_{(s)}$ + $2HCl_{(aq)}$ $\longrightarrow$ $ZnCl_{2(aq)}$ + $H_{2(g)}$

Number of particles:

1 atom	:	2 molecules	:	1 molecule	:	1 molecule
6.02×10^{23}	:	6.02×10^{23}	:	6.02×10^{23}	:	6.02×10^{23}

In terms of moles:

1 mole	:	2 moles	:	1 mole	:	1 mole

The reaction between zinc and hydrochloric acid forming carbon dioxide and water.

The above illustration shows us that the mole ratio is proportional to the number of atoms or molecules taking part in a chemical reaction. Therefore, we can convert the number of atoms or molecules taking part in a chemical reaction directly into moles.

The equation can be interpreted as 1 mole of zinc reacting with 2 moles of hydrochloric acid to produce 1 mole of zinc chloride and 1 mole of hydrogen.

The use of moles is very useful as it allows chemical calculations to be done and expressed as grams or kilograms and volumes of gases as cubic decimetre or cubic centimetre.
The following step by step worked examples show how chemical calculations are carried out.

Example 1:

Magnesium (Mg) reacts with dilute hydrochloric acid (HCl) according to the equation below:

magnesium + hydrochloric acid $\longrightarrow$ magnesium chloride + hydrogen gas

$$Mg_{(s)} + 2HCl_{(aq)} \longrightarrow MgCl_{2(aq)} + H_{2(g)}$$

In an experiment, 2g of magnesium ribbon is allowed to react with excess hydrochloric acid.

a. How many moles of magnesium have reacted?
b. What mass of magnesium chloride would be formed?
c. What is the volume of hydrogen at room temperature and pressure that would be produced?

Solution:

Step 1: Write down the balanced equation:

Magnesium + hydrochloric acid → magnesium chloride + hydrogen gas

$Mg_{(s)} + 2HCl_{(aq)} \rightarrow MgCl_{2(aq)} + H_{2(g)}$

Step 2: Write down the mole ratio.

1 mol Mg : 2 mol HCl → 1 mol $MgCl_2$: 1 mol H_2

1 mole of magnesium reacts with 2 moles of hydrochloric acid to produce 1 mol of magnesium chloride and 1 mole of hydrogen gas.

Step 3: Using the molar ratio, you can find the number of moles of magnesium.

a. Number of moles of magnesium (Mg) $= \frac{mass}{A_r}$

$= \frac{2}{24}$

$=$ **0.0833mol**

b. From the equation, 1 mol of magnesium (Mg) produces 1mol of magnesium chloride ($MgCl_2$). So, 0.833mol of magnesium (Mg) will produce 0.0833mol of magnesium chloride ($MgCl_2$).

Therefore, the mass of magnesium chloride ($MgCl_2$) produced is = 0.0833 x 95

= **7.91g**

Note: The total molar mass (M_r) of magnesium chloride ($MgCl_2$) is (1 x 24) + (2 x 35.5) = **95**

c. From the equation, 1 mol of magnesium (Mg) produces 1 mol of hydrogen gas (H_2).

So, 0.833 mol of magnesium (Mg) will produce 0.833 mol of hydrogen gas (H_2).

Volume of hydrogen gas (H_2) at r.t.p = number of mole (n) x 24dm^3

= 0.833 x 24

= **20dm^3**

Volume of hydrogen gas (H_2) at s.t.p = number of mole (n) x 22.4 dm^3

= 0.833 x 22.4

= 1.8dm^3

Example 2:
Sodium reacts with water according to the equation:

sodium	+	water	$\longrightarrow$	sodium hydroxide	+	hydrogen gas
$2Na_{(s)}$	+	$2H_2O_{(l)}$	$\longrightarrow$	$2NaOH_{(aq)}$	+	$H_{2(g)}$

a. How many moles of water are required to react with one mole of sodium?
b. What is the volume of hydrogen produced at r.t .p and s.t.p when 1 mole of sodium is reacted with water?

Solution:
Step 1: Write down the balanced equation.

sodium	+	water	$\rightarrow$	sodium hydroxide	+	hydrogen gas
$2Na_{(s)}$	+	$2H_2O_{(l)}$	$\rightarrow$	$2NaOH_{(aq)}$	+	$H_{2(g)}$

Step 2: Write down the mole ratio.

2mol Na : 2mol H_2O 2mol NaOH : 1 mol H_2

2 moles of sodium reacts with 2 moles of water to produce 2 moles of sodium hyroxide and 1 mole of hydrogen gas.

Step 3: Use the molar ratio to answer the questions.

a. From the equation, 2 mol of sodium (Na) react with 2 mol of water (H_2O)
Therefore, 1 mol of water (H_2O) will react with 1 mol of sodium (Na).

b. From the equation, 2 mol of sodium (Na) produces 1 mol of hydrogen gas (H_2).
So, 1 mol of sodium (Na) will produce 0.5 mol of hydrogen gas (H_2).

Volume of hydrogen gas (H_2) at r.t.p = number of mole (n) x $24dm^3$
= 0.5 x 24
= **$12dm^3$**

Volume of hydrogen gas (H_2) at s.t.p = number of mole (n) x $22.4dm^3$
= 0.5 x 22.4
= **$11.2dm^3$**

Now, check what you have just learnt by trying out the learning activity below!

Learning Activity 12 **60 minutes**

Answer the following questions:

Show your working out in the boxes provided.

1. a. What mass of magnesium oxide is produced when 6g of magnesium burns completely with oxygen?

   Given the relative atomic mass of magnesium (Mg) = 24 and oxygen (O) = 16.

   magnesium + oxygen $\longrightarrow$ magnesium oxide

   $2Mg_{(s)} + O_{2gl)} \longrightarrow 2MgO_{(s)}$

   b. What volume of oxygen gas (O_2) at r.t.p and s.t.p has reacted with 6g of magnesium?

2. a. A reaction occurs between sodium chloride (NaCl) solution and silver nitrate solution ($AgNO_3$).

   If we have 5.85g of sodium chloride (NaCl), how many grams silver nitrate ($AgNO_3$) would we need for all the sodium chloride (NaCl) to be used up, given the relative atomic mass of sodium (Na) = 23, chlorine (Cl) = 35.5, silver (Ag) = 108, and nitrogen (N) = 16?

   The equation for the reaction is:

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

   $NaCl_{(aq)} + AgNO_{3(aq)} \longrightarrow AgCl_{(s)} + NaNO_{3(aq)}$

b. Using the same reaction above, how much silver chloride (AgCl) will be produced?

The relative atomic mass of silver (Ag) = 108 and chlorine (Cl) = 35.5.

3. A reaction occurs between copper chloride ($CuCl_2$) and sodium hydroxide (NaOH) solution to produce a precipitate of copper hydroxide ($Cu(OH)_2$) in a solution of sodium chloride (NaCl).

The equation for the reaction is:

copper chloride	+ sodium hydroxide	$\longrightarrow$	copper hydroxide	+	sodium chloride
$CuCl_{2(aq)}$	+ $2NaOH_{(aq)}$	$\longrightarrow$	$Cu(OH)_{2(s)}$	+	$2NaCl_{(aq)}$
1 mole	2 moles		1 mole		2 moles

The molar masses:

135	40	98	58.5

If we wanted to make 10 g of copper hydroxide ($Cu(OH)_2$), what mass of copper chloride ($CuCl_2$) and sodium hydroxide (NaOH) should we start with? (Relative atomic mass of copper (Cu) = 64, chlorine (Cl) = 35.5, sodium (Na) = 23, oxygen (O) = 16 and hydrogen (H) = 1).

4. a. Calculate the mass of copper(Cu) produced by the complete reduction of 5.0g of copper(II) oxide (CuO) by hydrogen, given the relative atomic masses of copper = 64, oxygen (O) = 16 and hydrogen (H) = 1.

The equation for the reaction is:

Copper (II) oxide	+	hydrogen gas	$\longrightarrow$	copper	+ water
$CuO_{(s)}$	+	$H_{2(g)}$	$\longrightarrow$	$Cu_{(s)}$	+ $H_2O_{(l)}$

b. Calculate the volume of hydrogen gas at s.t.p. that reacts with copper(II)oxide(CuO).

5. Barium sulphate ($BaSO_4$) precipitates when sodium sulphate (Na_2SO_4) solution is added to barium chloride ($BaCl_2$) solution as shown in the equation below.

   Calculate the amount of barium chloride ($BaCl_2$) needed to prepare 50.0g of barium sulphate ($BaSO_4$).

   Relative atomic masses are barium (Ba) = 137, sulphur (S) = 32, oxygen (O) = 16, sodium (Na) = 23, chlorine (Cl) = 35.5

barium chloride	+	sodium sulphate	$\rightarrow$	barium sulphate	+	sodium chloride
$BaCl_{2(aq)}$	+	$Na_2SO_{4(aq)}$	$\rightarrow$	$BaSO_{4(s)}$	+	$2NaCl_{(aq)}$
1 mole		1 mole		1 mole		2 moles

Thank you for completing your learning activity 12. Check your work. Answers are at the end of this module.

Limiting and excess reactant

It is impossible to clean your entire home with a small amount of soap to remove all the dirt. As there is insufficient soap, chemists would say that soap is the limiting reactant.

For any chemical reaction, it is possible to calculate the exact quantities of reactants that are required and products that are formed from a balanced chemical equation. Reactions should be carried out using exact quantities of reactants to reduce wastage.

In some cases, there are reactions that are carried out using an excess amount of one reactant. This ensures that the more expensive reactant is completely used up. To do so, we make use of the idea of limiting reactants.

The reactant that is completely used up in a reaction is known as **limiting reactant.** It is called limiting reactant because it determines or limits the amount of products formed. The reactants that are not used up are called the **excess reactants**.

What limits the amount of products formed?

Consider the reaction between hydrogen gas (H_2) and chlorine gas (Cl_2) to form hydrogen chloride or hydrochloric acid (HCl).

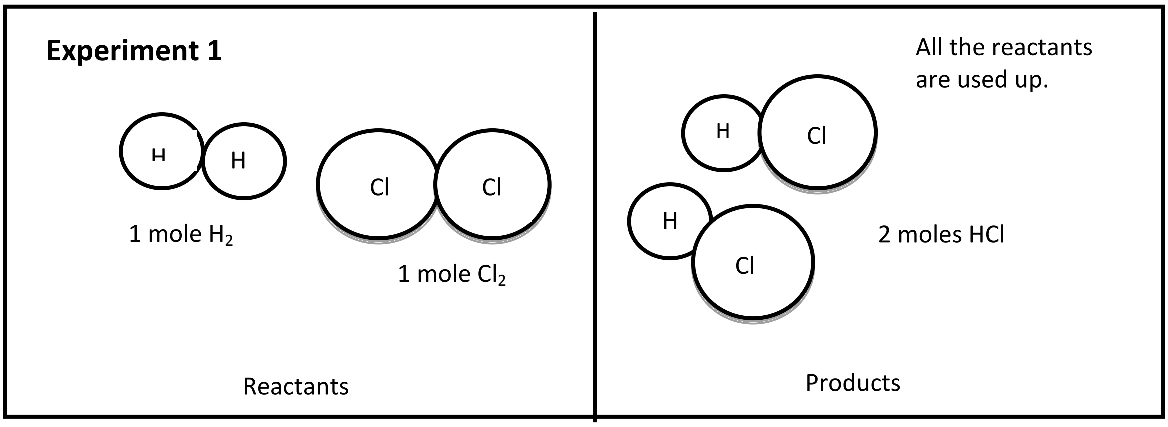
The chemical equation for the reaction is shown below.

hydrogen	+	chlorine	$\longrightarrow$	hydrogen chloride
$H_{2(g)}$	+	$Cl_{2(g)}$	$\longrightarrow$	$2HCl_{(aq)}$
1 mole		1 mole		2 moles

From the equation, we can see that one mole of hydrogen reacts with one mole of chlorine to produce two moles of hydrogen chloride.

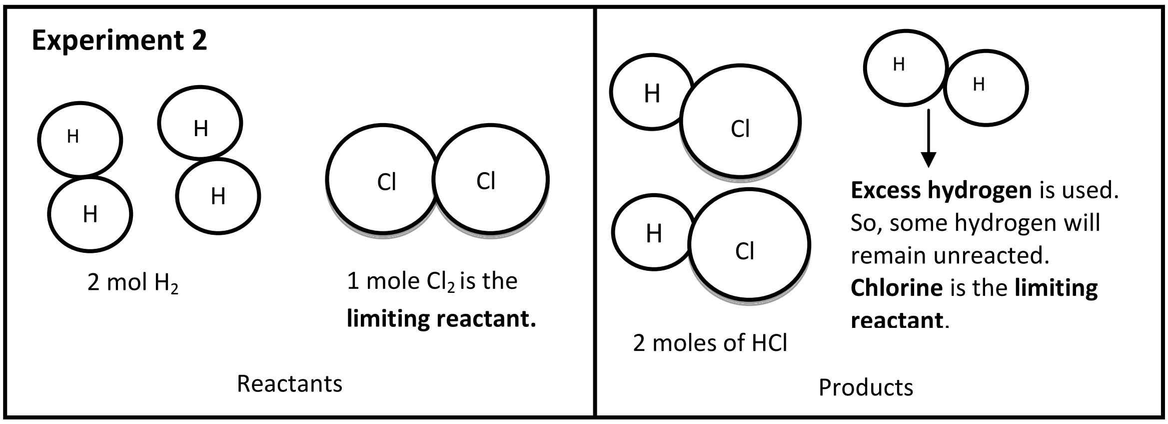
The diagram below shows the results of three experiments obtained by using different molar ratios of the reactants, hydrogen and chlorine.

In **experiment 1**, the ratio of the number of moles of reactant molecules used is the same as those given in the balanced chemical equation. One mole of hydrogen reacts with one mole of chlorine to give two moles of hydrogen chloride.

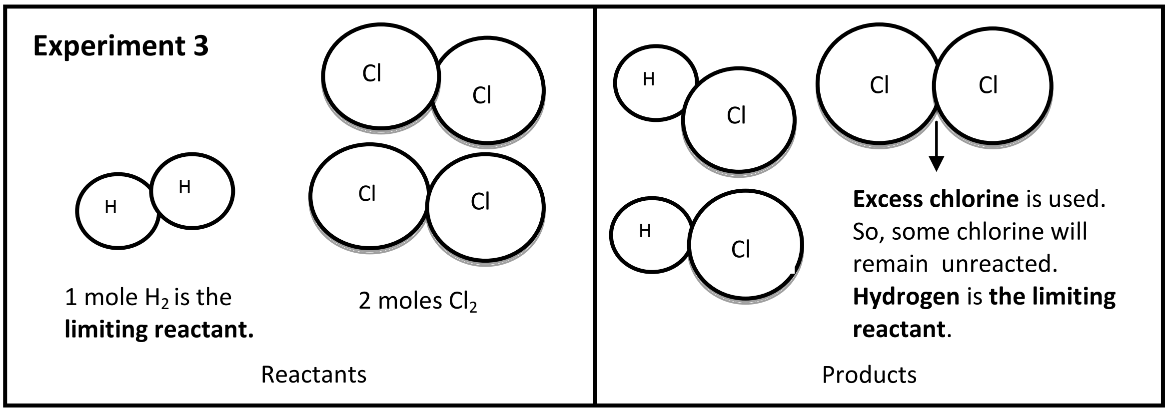


Experiment 1. A reaction between hydrogen and chlorine gas showing that they are all used up.

In experiments 2 and 3 below, the ratio of the number of moles of reactant molecules used is different from those given in the chemical equation.



Experiment 2. A reaction between hydrogen and chlorine showing the used of excess hydrogen. Chlorine is the limiting reactant.



Experiment 3. A reaction between hydrogen and chlorine showing the used of excess chlorine. Hydrogen is the limiting reactant.

The reactant that is completely used up is known as the limiting reactant. The reactants that are not used up are called the excess reactants.

Example 1:
Zinc reacts with hydrochloric acid according to the equation below:

zinc	+	hydrochloric acid	$\longrightarrow$	zinc chloride	+	hydrogen
$Zn_{(s)}$	+	$2HCl_{(aq)}$	$\longrightarrow$	$ZnCl_{2(aq)}$	+	$H_{2(g)}$
1 mol		2 mol		1 mol		1 mol

If 0.05 mol of zinc is added to 0.075 mol of hydrochloric acid,

a. identify the limiting reactant.
b. calculate the amount in moles of the excess reactant that will remain unreacted.

Solution:

a. According to the equation, 1 mol of zinc (Zn) reacts with 2 mol of hydrochloric acid (HCl), therefore 0.05 mol of zinc (Zn) reacts with 0.05 x 2 = 1 mol of hydrochloric acid (HCl).

   0.075 mol of hydrochloric acid is used and this will react with 0.075/2 = 0.0375 mol of zinc. Since 0.05 mol of zinc is used, the zinc must be in excess and **hydrochloric acid** is the **limiting reactant**.

b. The amount of zinc which remains unreacted is:
   = 0.05 – 0.0375
   = **0.0125 mol**

Example 2:
Ethene (C_2H_4) burns in oxygen (O_2) to form carbon dioxide (CO_2) and water (H_2O) according to the equation below:

ethene	+	oxygen	$\longrightarrow$	carbon dioxide	+	water
$C_2H_{4(g)}$	+	$3O_{2(g)}$	$\longrightarrow$	$2CO_{2(g)}$	+	$H_2O_{(l)}$
1 mol		3 mol		2 mol		1 mol
1 volume		3 volumes		2 volume		1 volume

In an experiment, $10cm^3$ of ethane (C_2H_4) is burnt in $50cm^3$ of oxygen (O_2).

(i) Which gas is supplied in excess?
(ii) Calculate the volume of the excess gas remaining at the end of reaction.
(iii) Calculate the volume of carbon dioxide produced.

Solution:
(i) According to the equation, 1 volume of ethane reacts with 3 volumes of oxygen.

Therefore, the volume of oxygen that is used;

$$= \frac{3 \times 10}{1}$$

$$= \mathbf{30cm^3}$$

However, $50cm^3$ of oxygen is used, therefore, **oxygen gas** is in **excess**.

(ii) The volume of oxygen remaining = initial volume − volume used

$= 50cm^3 - 30cm^3$

$= \mathbf{20\ cm^3}$

(iii) 1 volume ethane (C_2H_4) produces 2 volumes of carbon dioxide (CO_2).

Therefore, volume of carbon dioxide produced $= \frac{2 \times 10}{1}$

$= \mathbf{20cm^3}$

Why is it important to identify the limiting reactant?
In the chemical industry, large amounts of chemicals are required to manufacture a particular product. To get the maximum yield of a product at the minimum cost, we need to know the limiting reactant.

We will choose the most expensive reactant to be the limiting reactant, and use excess amounts of the other reactants in a reaction. This ensures, the expensive reactant is used up. In many industrial reactions, excess reactants are recycled as far as possible in order to reduce the production costs.

Now, check what you have just learnt by trying out the learning activity below!

Learning Activity 13 **60 minutes**

Answer the following question:

Show your working out in the boxes provided.

1. A solution containing 20g of sodium chloride (NaCl) is reacted with a solution containing 20g of silver nitrate according to the equation given below.

   The relative atomic masses are: sodium (Na) = 23, chlorine (Cl) = 35.5, silver (Ag) = 108, nitrogen (N) = 14, and oxygen (O) = 16.

sodium chloride	+	silver nitrate	$\longrightarrow$	silver chloride	+	sodium nitrate
$NaCl_{(aq)}$	+	$AgNO_{3(aq)}$	$\longrightarrow$	$AgCl_{(s)}$	+	$NaNO_{3(aq)}$
1mol		1mol		1mol		1mol

   Molar masses:

58.5		170		143.5		85

   a. Which is the limiting reactant and which is the excess reactant?

   b. What mass of silver chloride will precipitate?

c. What mass of excess reactant will remain unreacted at the end of the reaction?

2. In an experiment 20g of zinc (Zn) powder is added to a solution containing 64g of copper sulphate ($CuSO_4$) according to the equation below.

The relative atomic masses are: Zinc (Zn) = 65, copper (Cu) = 64, sulphur (S) = 32, and oxygen (O) = 16.

zinc	+	copper sulphate	$\longrightarrow$	zinc sulphate	+	copper
$Zn_{(s)}$	+	$CuSO_{4(aq)}$	$\longrightarrow$	$ZnSO_{4(aq)}$	+	$Cu_{(s)}$
1mol		1mol		1mol		1mol

Molar masses:

65	160	161	64

a. Which reactant is in excess? Which is the limiting reactant?

b. Calculate the mass of copper (Cu) produced.

c. Calculate the mass of zinc sulphate ($ZnSO_4$) produced.

d. What mass of excess reactant will remain unreacted at the end of the reaction?

Thank you for completing your learning activity 13. Check your work. Answers are at the end of this module.

Percentage and theoretical yield

Calculations based on chemical equations give the theoretical yield of product to be expected from a reaction. **Theoretical yields** are the amount of products calculated from the complete reaction of the limiting reactant. The **actual yields** are the amount of the products that are actually produced in a reaction.

Often, the actual yield is almost always **less** than the calculated yield of the product. The reason may be that some products have remained in a solution or on a filter paper and has not been weighed with the final yield.

The percentage yield of a product is shown in a relationship below:

$$\text{Percentage yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100\%$$

Example 1:
When 1.92g of magnesium (Mg) is heated in excess oxygen (O), 3.0g of magnesium oxide (MgO) is obtained.

The relative atomic masses are: Mg = 24, O = 16.

Calculate the percentage yield of magnesium oxide (MgO).

Solution:
Step 1 Write a balanced equation for the reaction.

magnesium + oxygen $\longrightarrow$ magnesium oxide

$$\mathbf{2Mg_{(s)} + O_{2(g)} \longrightarrow 2MgO_{(s)}}$$

Molar ratios:

2 moles 1 mole 2 moles

Step 2 Use the formula below to find the number of moles of 1.92g of magnesium (Mg) given.

The formula is: $\mathbf{n = \frac{m}{A_r}}$

This means that 24g of magnesium will produce: $= \frac{1.92}{24}$

$= \mathbf{0.08mol}$

Step 3 Use the number of moles magnesium (0.08 mol) to multiply with the molar mass of magnesium oxide (40g) to calculate for the mass of magnesium oxide and use the relationship below:

$$\mathbf{m = n \times Mr}$$

Therefore, the mass of magnesium oxide (MgO) using 0.08mol of magnesium is:

$$m = 0.08 \times 40$$
$$= \mathbf{3.2g}$$

This means, that 24 g of magnesium (Mg) should produce; $= \dfrac{1.92 \times 40}{24}$

$$= \mathbf{3.2\ g}$$

Step 4 Find the percentage yield of magnesium oxide (MgO) using the relationship below:

$$\text{Percentage yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100\%$$

From the question given in the previous page, the actual amount of magnesium oxide (MgO) obtained was 3.0g.

Therefore, the percentage yield of magnesium oxide (MgO) is:

$$\text{Percentage yield} = \frac{3.0}{3.2} \times 100\%$$
$$= \mathbf{93.8\%}$$

Example 2:
Nitrogen (N_2) and hydrogen (H_2) react to form ammonia (NH_3). A hydrogen gas with a volume of 12dm^3 reacts with excess nitrogen to form 2dm^3 of ammonia.

What is the percentage yield of ammonia at room temperature and pressure? (1 mol of gas occupies 24 dm^3 at room temperature and pressure).

Solution:
Step 1 Write a balanced equation for the reaction.

nitrogen + hydrogen $\longrightarrow$ ammonia

$$N_{2(g)} + 3H_{2(g)} \longrightarrow 2NH_{3(g)}$$

Molar ratios:

1 mole 3 moles 2 moles

Step 2 Use the formula below to find the number of moles of hydrogen (H_2).

The formula is: $$n = \frac{V}{24dm^3 \text{ at r.t.p.}}$$

From the equation given, 3 moles of hydrogen(H_2) reacted with 1 mole of nitrogen(N_2) produced 2 moles of ammonia(NH_3).

Therefore, the number of moles of hydrogen used is $= \frac{12}{24}$

$= $ **0.5mol**

Step 3 Find the molar ratio of hydrogen (H_2) and ammonia (NH_3) to find the **theoretical moles** of ammonia (NH_3).

From the equation, the molar ratio of hydrogen (H_2) and ammonia (NH_3) is **3:2 ratios.**

Therefore, the theoretical moles moles of ammonia (NH_3) formed is:

$$= \frac{0.5 \times 2}{3}$$

$$= \mathbf{0.33\ mol}$$

Step 4 Find the **theoretical volume** of ammonia (NH_3) using the relationship below:

$$\mathbf{V(r.t.p.) = n \times 24}$$

From the question given, the **actual volume** of ammonia (NH_3) formed is **2dm^3**.

Therefore, the theoretical volume of ammonia (NH_3) is:

$$= 0.33 \times 24$$

$$= \mathbf{8.0\ dm^3}$$

Hence, the percentage yield of ammonia (NH_3) is:

$$= \frac{2}{8.0} \times 100\%$$

$$= \mathbf{25\%}$$

Example 3:
A student calculates that a certain reaction will yield 7.0g of a salt. Her product weighs 6.3g. What percentage yield has she obtained?

Solution:

$$\text{Percentage yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100\%$$

$$\text{Percentage yield} = \frac{6.3}{7.0} \times 100\%$$

$$= \mathbf{90\ \%}$$

Now, check what you have just learnt by trying out the learning activity below!

Learning Activity 14 **80 minutes**

Answer the following questions:

Show your working out in the boxes provided.

1. When 6.4g of copper (Cu) were heated in air, 7.6g of copper(II) oxide(CuO) were obtained.
   The relative atomic masses are: Copper (Cu) = 64 and oxygen (O) = 16.

   a. Calculate the mass of copper(II) oxide(CuO) that would be formed if the copper (Cu) reacted completely.

   b. Calculate the percentage yield of copper(II) oxide(CuO) that was actually obtained.

2. 28g of nitrogen (N_2) is reacted with hydrogen (H_2), 3.4g of ammonia (NH_3) has formed. The relative atomic masses are: Nitrogen (N) = 14 and hydrogen (H) = 1.

The equation for the reaction is shown below.

$$\begin{array}{cccc} \text{nitrogen} & + \text{ hydrogen} & \longrightarrow & \text{ammonia} \\ N_{2(g)} & + \ 3H_{2(g)} & \longrightarrow & 2NH_{3(g)} \end{array}$$

Molar ratios:

1 mole   3 moles   2 moles

a. What is the maximum mass of ammonia (NH_3) that could be formed?

b. What percentage of this yield was obtained?

3. When 98g of sulphuric acid (H_2SO_4) reacted with ammonia (NH_3), 120g of ammonium sulphate $(NH_4)_2SO_4$ are obtained.

The relative atomic masses are: hydrogen (H) = 1, sulphur (S) = 32, oxygen (O) = 16 and nitrogen (N) = 14.

The balanced equation for the reaction is:

$$\begin{array}{ccccc} \text{sulphuric acid} & + & \text{ammonia} & \rightarrow & \text{ammonium sulphate} \\ H_2SO_4 & + & 2NH_3 & \rightarrow & (NH_4)_2SO_4 \end{array}$$

a. Calculate the theoretical yield of ammonium sulphate $(NH_4)_2SO_4$.

b. Calculate the actual percentage yield.

4. 28.0g of nitrogen (N_2) reacted with 8.0g of hydrogen (H_2) to form 5.1g of ammonia (NH_3).

Relative atomic masses are: Nitrogen (N) = 14 and hydrogen (H) = 1.

a. Write a balanced equation for the reaction.

b. Which reactant is excess and which one is the limiting reactant?

c. What is the percentage yield of ammonia (NH_3)?

Thank you for completing your learning activity 14. Check your work. Answers are at the end of this module.

12.1.6 Solutions

Concentration of standard solutions in molarity

Standard solution is a solution containing known concentration of an element or a substance and a known weight of solute that is dissolved to make a specific volume. It is prepared using a standard solution called **primary standard**.

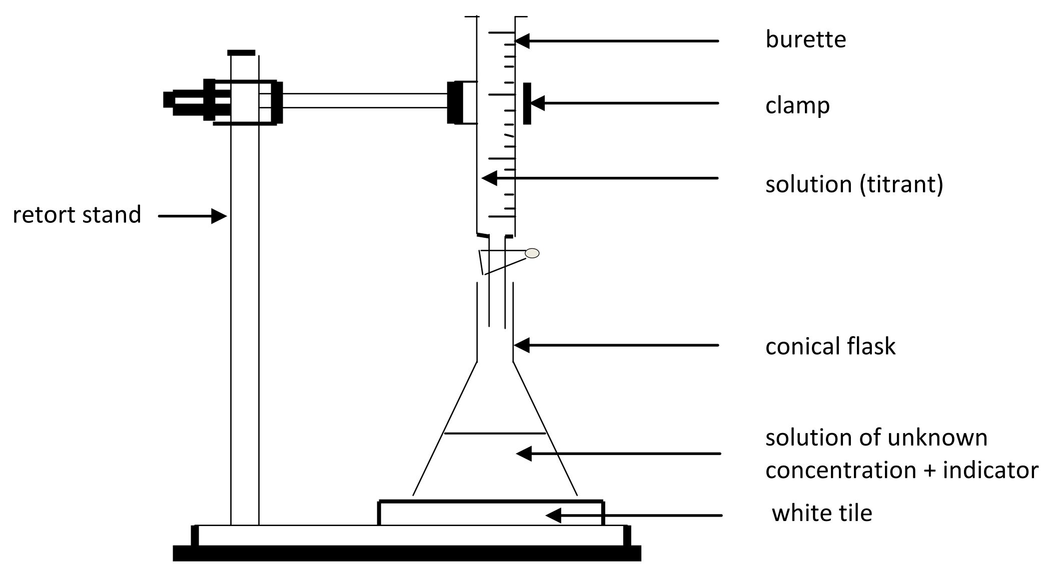
Standard solutions are used to determine the concentrations of other substances, such as solutions in titrations. The concentrations of standard solutions are normally expressed in units of moles per liter or moles per cubic decimetre. The unit **mol/L and mol/dm^3** are often abbreviated to **M** for **molarity** (another term for concentration).

A standard solution can be calculated by reacting it against a solution of alkali of known concentration. The method used is called **titration.** Once this has been calculated, it can then be used as a standard solution to find the concentration of a solution of an alkali.

The solution with known concentration is called a standard solution. The concentration of such a solution is usually expressed in molarity, that is, moles per cubic decimetre (mol/dm^3).

The unit mol/L and mol/dm^3 are often abbreviated to M for molarity (another term for concentration).

APPARATUS USED FOR TITRATION EXPERIMENTS



From our previous lesson we have already learnt the formulas given below. We will still use these formulas to calculate the concentration, number of moles and volumes of unknown solutions.

It is important to remember that in calculating concentration of solution, the volumes should be in cubic decimetre (dm^3) or in litres (L). Any other units of volume must be converted to cubic decimetre (dm^3) of litres (L).

For example: To convert $50cm^3$ to dm^3, divide it in 1000.

$$= \frac{50cm^3}{1000}$$

$$= \mathbf{0.05dm^3}$$

$$\text{Concentration(C) in mol/dm}^3 = \frac{\text{Number of moles(n) in mol}}{\text{Volume of solution(V) in dm}^3}$$

$$\text{Number of moles(n) in mol} = \text{concentration(C) in mol/dm}^3 \times \text{Volume(V) in dm}^3$$

$$\text{Volume(V) in dm}^3 = \frac{\text{Number of moles(n) in mol}}{\text{Concentration(C) in mol/dm}^3}$$

Let us do some more examples on concentration, number of moles and volumes of solutions:

Example

A household ammonia solution (NH_3) was analysed to determine its ammonia content. $25cm^3$ of ammonia solution needed $21.9cm^3$ of $0.11mol/dm^3$ sulphuric acid (H_2SO_4) to reach the end-point of titration.

Calculate the concentration of ammonia (in g/dm^3), in the household ammonia solution. The equation for the reaction is shown below:

sulphuric acid	+	ammonia	→	ammonium sulphate
$H_2SO_{4(aq)}$	+	$2NH_{3(g)}$	→	$(NH_4)_2SO_{4(s)}$
1mol		2mol		1mol

$C = 0.11mol/dm^3$ $V = 25cm^3$
$V = 21.9cm^3$ $C = ?$
$n = ?$

(i) Use the formula below to calculate for the number of moles of sulphuric acid.

Number of moles of sulphuric acid (H_2SO_4) $= C \times V$

$$= \frac{0.11 \times 21.9}{1000}$$

$$= 0.11 \times 0.0219$$

$$= \mathbf{2.409 \times 10^{-3} mol}$$

(ii) Find the number of moles of ammonia (NH_3).

From the equation:
1 mol of sulphuric acid (H_2SO_4) reacts with 2 mol of ammonia (NH_3) solution.

The number of moles of ammonia in $25cm^3$ $= 2 \times 2.409 \times 10^{-3} mol$

$$= \mathbf{4.818 \times 10^{-3} mol}$$

The number of moles of ammonia in $1000cm^3$ $= 4.818 \times 10^{-3} \times \frac{1000}{25}$

$$= \mathbf{0.1927\ mol}$$

M_r of $NH_3 = (1 \times 14) + (3 \times 1) = 17$

Hence, the concentration of ammonia (NH_3) in g/dm^3 is :

$$= 0.1927 \times 17$$

$$= \mathbf{3.28} \frac{g}{dm^3}$$

Now, check what you have just learnt by trying out the learning activity below!

Learning Activity 15 **60 minutes**

Answer the following questions:

1. $25cm^3$ of a $0.100mol/dm^3$ sodium hydroxide solution was titrated with $0.075mol/dm^3$ hydrochloric acid.

a. Write a balanced equation for the reaction.

____

b. How many cubic decimeters is $25cm^3$? ($1000cm^3 = 1dm^3$). ___________

c. Calculate the number of moles of sodium hydroxide that was used to react with 0.075 mol/dm^3 hydrochloric acid.

Thank you for completing your learning activity 15. Check your work. Answers are at the end of this module.

Dilution

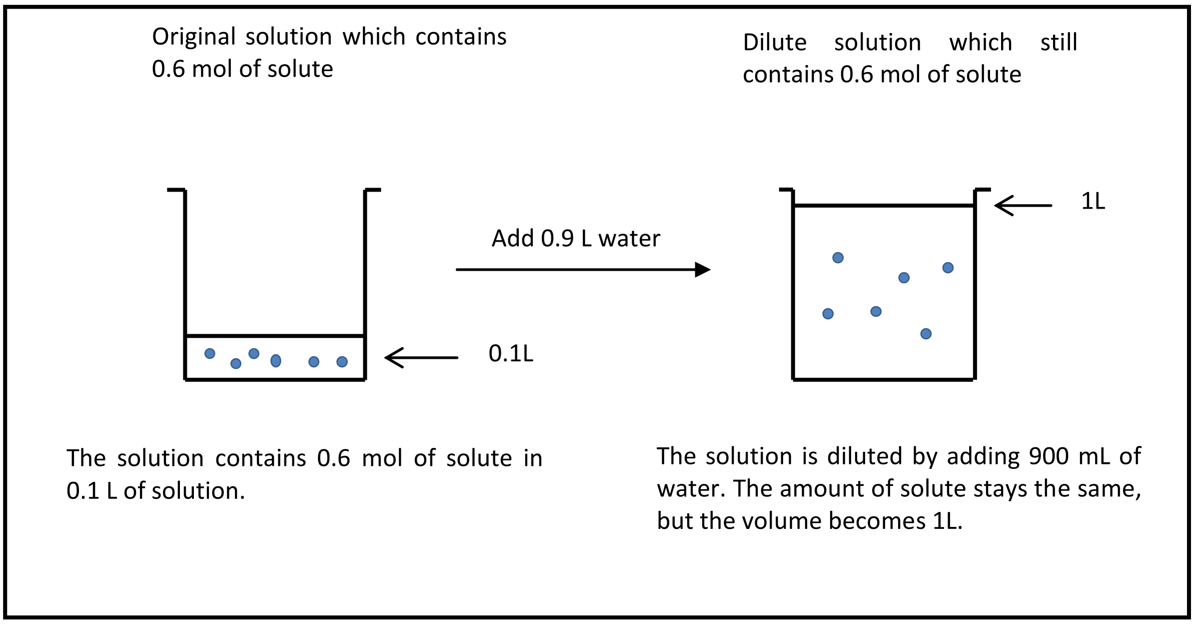
An aqueous solution with known concentration is diluted (weakened) by adding more water to it. Therefore, **dilution** is the process of weakening the concentration of a solute in solution by simply mixing with more solvent.

When this happens the number of moles of solute remains the same. If additional solute is added or some water is removed (by evaporation), a solution becomes more concentrated.

For example, you can add water to a concentrated orange juice to dilute it until it reaches a concentration that is pleasant to drink.

Number of moles before dilution = number of moles after dilution

The diagram below shows the dilution of a concentrated solution:



The dilution of solution showing the addition of more volume of water but the amount of solute remains the same.

If we calculate the concentrations of the two solutions given in the previous page, they will have the different concentrations.

Original solution	Diluted solution
$c = \frac{n}{V}$ $= \frac{0.6\text{mol}}{0.1\text{L}}$ $=$ **6 mol/L**	$c = \frac{n}{V}$ $= \frac{0.6\text{mol}}{1.0\text{L}}$ **=** 0.6mol/L

It shows that increasing the volume of water will decrease the concentration of solution.

The formula to use for diluting a solution is shown below:

$$\boxed{\mathbf{C_1V_1 = C_2V_2}}$$

Where:

C_1 = initial concentration of solution
V_1 = initial volume of solution
C_2 = final concentration of solution
V_2 = final volume of solution

Example 1:
If 200cm^3 (0.2dm^3) of a 0.4M (mol/dm^3) detergent solution is diluted to 800cm^3 (0.8dm^3), what is the new concentration?

Solution:

C_1 = 0.4 M
V_1 = 200cm^3(0.2dm^3)
C_2 = ?
V_2 = 800cm^3(0.8 dm^3)

Final concentration is unknown. Therefore, the formula, $C_1V_1 = C_2V_2$ is rearranged;

$$C_2 = \frac{C_1V_1}{V_2}$$

$$= \frac{(0.4\text{mol/dm}^3)(0.2\text{dm}^3)}{0.8\text{dm}^3}$$

$$= \mathbf{0.1\ mol/dm^3\ or\ 0.1\ M}$$

It is important to remember that in calculating concentration of solution, the volumes should be in cubic decimetre (dm^3) or in litres (L). Any other units of volume must be converted to cubic decimetre (dm^3) of litres (L).

Example 2:
Calculate the concentration of $25cm^3$ ($0.025dm^3$) of a 0.100M sugar solution when it is diluted to $125cm^3$ ($0.125dm^3$).

Solution:

C_1 = 0.10 M
V_1 = 25 cm^3 (0.025 dm^3)
C_2 = ?
V_2 = 125 cm^3 (0.125 dm^3)

Final concentration is unknown. Therefore, the formula, $C_1V_1 = C_2V_2$ is rearranged;

$$C_2 = \frac{C_1V_1}{V_2}$$

$$= \frac{(0.10mol/dm^3)(0.125dm^3)}{0.025dm^3}$$

= **0.5 mol/dm³ or 0.5 M**

Now, check what you have just learnt by trying out the learning activity below!

Learning Activity 16 **60 minutes**

Answer all the questions. Write your working out in the boxes provided.

1. Calculate the concentration of $20cm^3$ ($0.2dm^3$) of a 0.20M salt solution when it is diluted to $250cm^3$ ($0.250dm^3$).

2. A bottle of concentrated hydrochloric acid (HCl) has 12.5M written on its label. A student takes $10.0cm^3$ ($0.01dm^3$) of this solution.

   To what volume of water must it be added in order for the student to get a 0.20 M acid solution?

3. 0.1 mol sodium hydroxide ($NaOH$) in 250mL (0.250L) is diluted to 1000mL (1L).

   Find the concentration of the solution before and after dilution.

Thank you for completing your learning activity 16. Check your work. Answers are at the end of this module.

REVISE WELL USING THE MAIN POINTS ON THE NEXT PAGE.

SUMMARY

You will now revise this module before doing Assessment 6. Here are the main points to help you revise. Refer to the module topic if you need more information.

- **Isotopes** are atoms of the same element which have the same number of protons but different number of neutrons.
    - For example: Chlorine exists naturally as a mixture of two isotopes: chlorine -35 (75%) and chlorine -37 (25%).

      The calculation below explains why the mass number of chlorine is not a whole number and is always given as 35 (0.75 x 35 + 0.25 x 37 = 35).

- **Relative atomic mass (A_r) of an element** is the average mass of one atom of the element compared to $^1/_{12}$ of the mass of an atom of carbon – 12.

$$\text{Relative atomic mass (Ar)} = \frac{\text{Average mass of one atom of an element}}{\text{Mass of } \frac{1}{12} \text{ of one atom of carbon-12}}$$

For example: The relative atomic mass of sodium (Na) atom is **23** and magnesium (Mg) atom is **24**.

- The symbol for **relative atomic mass** is $\mathbf{A_r}$.

- **Relative molecular mass (M_r)** of a substance is the average mass of a molecule of the substance when compared to $^1/_{12}$ of the mass of an atom of carbon -12.

$$\text{Relative molecular mass (Ar)} = \frac{\text{Average mass of one molecule of a substance}}{\text{Mass of } \frac{1}{12} \text{ of one atom of carbon-12}}$$

- The symbol **for relative molecular mass** is $\mathbf{M_r}$.
- The **relative molecular mass** of a substance is equal to the sum of the relative atomic masses of the atoms in the formula.

For example; the relative molecular mass of water (H_2O) is (1x2) + (1 x 16) = **18**.

- **Relative formula** mass is used for ionic compounds and is equal to the sum of all relative atomic masses of the atoms in the formula.
  - For example; the relative formula mass of sodium chloride (NaCl) is (1 x 23) + (1 x 35.5) = 58.5.

- One mole of particles contains **6.02×10^{23}**. This number is called **Avogadro's number**.
- The formulas to use in moles and number of particles are:
  - $\text{Number of moles} = \dfrac{\text{number of particles}}{\text{Avogadro's number}}$
  - Number of particles = amount of substance in mol x Avogadro's number

- **Molar mass (M_r or M)** the mass of one mole of particles of any substance. The unit used is **grams per mole (g/mol).**

- The formulas to use in moles and molar masses are:
  - $\text{Number of moles} = \dfrac{\text{mass of element in grams (g)}}{\text{relative atomic mass of element}}$
  - $\text{Number of moles} = \dfrac{\text{mass of element in grams (g)}}{\text{relative molecular mass of a substance}}$
  - $\text{Number of moles} = \dfrac{\text{mass of element in grams (g)}}{\text{relative formula mass of ionic compound}}$

- **Percentage composition of a compound** – tells us the percentage of each element present in the formula of the compound.

$$\text{Percentage by mass of an element} = \frac{A_r \text{ x number of atoms in formula}}{M_r \text{ of a compound}} \times 100\%$$

For example; the percentage of oxygen (O) atom in water:

One mole of water is (1 x 2) + (1 x 16) = 18. The molar mass of water is **18g/mol**. Therefore, the percentage of oxygen (O) atom is:

$$\% \text{ of oxygen (O)} = \frac{\text{Relative atomic mass of oxygen (O)}}{\text{Molar mass of water } (H_2O)} \times 100\ \%$$

$$= \frac{16}{18}$$

$$= \mathbf{88.8\ \%}$$

- Equal volumes of all gases under the same temperature and pressure contain the same number of molecules. This allows us to express the stoichiometry of the equation for reacting gases in terms of their volumes.

- temperature and pressure (r.t.p.), the volume of any gas is **24 dm^3 or 24,000 cm^3**.
- At standard temperature and pressure (s.t.p), the volume of any gas is 22.4 dm^3 or 22,400 cm^3.

- The formulas to use in moles and volume of any gas are:

$$\text{Number of moles gas} = \frac{\text{mass of gas in grams}}{M_r \text{ of gas}}$$

$$\text{Number of moles gas} = \frac{\text{volume of gas at r.t.p.}}{24 dm^3 \text{ or } 24{,}000 cm^3}$$

Volume of gas (r.t.p.) = number of moles x 24dm^3 or 24,000cm^3

Volume of gas (s.t.p.) = number of moles x 22.4dm^3 or 22,400cm^3

- **Empirical formula** of a compound can be found from its percentage composition. The empirical formula is the formula which only shows the simplest ratio of the elements present in a molecule.
- **Molecular formula** is the true formula which shows all the atoms present in a molecule. It can be calculated from the empirical formula and the relative molecular mass.
- **Stoichiometry (stoi-kio-me-tree)** is the calculation of relative quantities of reactants and products in chemical reactions. Stoichiometry is founded on the law of conservation of mass where the total mass of the reactants are equal to the total mass of the products.
- **Limiting reactant** is the reactant that is completely used up in a chemical reaction. It is called limiting reactant because it determines or limits the amount of products formed.
- **Excess reactant** – is the reactant that is not completely used up in a chemical reaction.
- The percentage yield of a product is shown in a relationship below:

$$\text{Percentage yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100\%$$

- **Theoretical yield of the product** is the calculation based on chemical equations.

- The **actual yield** is often less than the theoretical yield of the product.
- The concentration of a solution is the mass of solute in $1dm^3$ of the solution.
- The concentration of a solution can be measured in grams per cubic decimetre (g/dm^3) or mole per cubic decimetre (mol/dm^3).
- The formulas for concentration (molarity) are as follows:

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

$$\text{Concentration in } g/dm^3 = \frac{\text{number of moles of solute}}{\text{volume of solution in } dm^3}$$

Mass of solute in grams = Volume of solution in dm^3 x Concentration in g/dm^3

Number of moles of solute = Volume of solution in dm^3 x Concentration in g/dm^3

- **Standard solution** is a solution containing known concentration of an element or a substance and a known weight of solute that is dissolved to make a specific volume.
- **Dilution** is the addition of more water in a solution with known concentration.

The formula for dilution is: $\mathbf{C_1V_1 = C_2V_2}$

Where C_1 and V_1 ate the initial concentration and volume of the solution while C_2 and V_2 are the final concentration and volume of the solution.

NOW YOU MUST COMPLETE ASSESSMENT 1 AND RETURN IT TO THE PROVINCIAL CENTRE CO-ORDINATOR.

Answers to Learning Activities 1-16

Learning Activity 1

1. Carbon -12 : 6 protons, 6 electrons and 6 neutrons.
   Carbon -13 : 6 protons, 6 electrons and 7 neutrons.

2. $^{56}_{27}Co$

   $^{58}_{27}Co$

   $^{60}_{27}Co$

3. $$Ar(Li) = \frac{(6.02 \times 7.4) + (7.02 \times 92.6)}{100}$$
   $$= \frac{44.548 + 650.052}{100}$$
   $$= \mathbf{6.946}$$

4. $$= \frac{92.2 \times 27.98}{100} + \frac{4.7 \times 28.98}{100} + \frac{3.1 \times 29.97}{100}$$
   $$= 25.80 + 1.36 + 0.92$$
   $$= \mathbf{28.08}$$

5. a. $^{13}_{6}C$ b. $^{39}_{19}K$ c. $^{23}_{11}Na$

6. **Isotopes** are atoms of the same element with the same number of protons but different number of neutrons.

Learning Activity 2

1. Molar mass of bromine (Br) is 79.9041 g/mol.

   79.9041 = X% of (78.9183) + (1 - X%) of (80.9163)
   79.9041 = 78.9183X + 80.9163 – 80.9163X
   1.9980X = 1.0122
   X = 0.50661 X 100%
   X% = 50.661% for Bromine -79

100% - 50.661% = 49.339% for Bromine -81
Percentage abundance for **Bromine - 79 is 50.661%** and **Bromine -81 is 49.339%.**

2. Molar mass of lithium (Li) is 6.941 g/mol.

6.941 = x% of (6.0151) + (100% - x%) of (7.0160)
6.941 = 6.0151x + 7.0160 – 7.0160x
1.0009X = 0.075
x = 0.075/1.0009 X 100%
x% = 7.49% lithium-6
100% - 7.49% = 92.51% for lithium-7

Percentage abundance for **Lithium-6 is 7.49%** and for **Lithium-7 is 92.51%.**

Learning Activity 3

1. a. sodium atoms = (2.0) (6.02×10^{23})
= $\mathbf{1.204 \times 10^{24}}$

b. nitrogen molecules = (0.10) (6.02×10^{23})
= $\mathbf{6.02 \times 10^{22}}$

c. carbon atoms = (20.0) (6.02×10^{23})
= $\mathbf{1.204 \times 10^{25}}$

d. water molecules = (4.2) (6.02×10^{23})
= $\mathbf{2.5 \times 10^{24}}$

2 a. iron atoms = $(1.0 \times 10^{-2})(6.02 \times 10^{23})$
= $\mathbf{6.02 \times 10^{21}}$

b. carbon dioxide molecules = $(4.62 \times 10^{-5})(6.02 \times 10^{23})$
= $\mathbf{2.78 \times 10^{19}}$

c. silicon atoms = $(1.6 \times 10^{-8})(6.02 \times 10^{23})$
= $\mathbf{9.6 \times 10^{15}}$

3 a. = (0.4) (2)
= **0.8mol**

b. = (1.2) (4)
= **4.8mol**

Learning Activity 4

1. The **relative molecular mass** of a molecule is the average mass of the substance when compared ½ of the mass of an atom of carbon -12.

   The **relative molecular mass** of a substance is equal to the sum of the relative atomic masses of the atoms in the formula.

   Any from the two answers are correct.

2. a. $M_r(HCl)$ = (1 x 1) + (1 x 35.5)
   = **36.5**

   b. $M_r(CH_4)$ = (1 x 12) + (1 x 4)
   = **16**

   c. $M_r(C_{12}H_{22}O_{11})$ = (12 x 12) + (1 x 22) + (11 x 16)
   = **342**

   d. $M_r(SO_2)$ = (1 x 32) + (2 x 16)
   = **64**

   e. $M_r(H_2SO_4)$ = (1 x 2) + (1 x 32) + (4 x 16)
   = **98**

3. a. $M_r(Ca(OH)_2)$ = (1 x 40) + (2 x 16) + (2 x 1)
   = **74**

   b. $M_r(KF)$ = (1 x 39) + (1 x 19)
   = **58**

   c. $M_r(CuO)$ = (1 x 64) + (1 x 16)
   = **80**

   d. $M_r(K_2S)$ = (2 x 39) + (1 x 32)
   = **110**

   e. $M_r(MgO)$ = (1 x 24) + (1 x 16)
   = **40**

Learning Activity 5

1. a. $m(H_2)$ = (1) (2)
   = **2g**

   b. $m(O_2)$ = (2) (16)
   = **32g**

   c. $m(N_2)$ = (2) (14)
   = **28g**

   d. $m(Cl_2)$ = (2) (35.5)
   = **71g**

2. a. $n(O_2) = \frac{32}{16}$
   = **2 mol**

   b. $n(N_2) = \frac{7}{14}$
   = **0.5mol**

   c. $n(H_2) = \frac{8}{1}$
   = **8mol**

   d. $n(Cl_2) = \frac{71}{35.5}$
   = **2mol**

3. a. $m(N_2)$ = (7) (14)
   = **98g**

   b. $m(H_2O)$ = (3) (1 x 2 + 1 x 16)
   = **54g**

   c. $m(HCl)$ = (4) (1 x 35.5)
   = **142g**

   d. $m(SO_2)$ = (0.5) (1 x32 + 2 x 16)
   = **32g**

Learning Activity 6

1. The compound is iron (III) oxide (Fe_2O_3).
   Iron (Fe) = (2 x 56) = 112
   Oxygen (O) = (3 x 16) = 48
   RMM of iron (III) oxide (Fe_2O_3) = 160

   $$\text{Percentage of iron (Fe)} = \frac{112}{160} \times 100\%$$
   $$= \mathbf{70\%}$$

2. The compound is potassium chloride (KCl).
   Potassium (K) = (1 x 39) = 39
   Chlorine (Cl) = (1 x 35.5) = 35.5
   RMM of potassium chloride (KCl) = 74.5

   $$\text{Percentage of potassium (K)} = \frac{39}{74.5} \times 100\%$$
   $$= \mathbf{52.3\%}$$

   Potassium nitrate (KNO_3)
   Potassium (K) = (1 x 39) = 39
   Nitrogen (N) = (1 x 14) = 14
   Oxygen (O) = (3 x 16) = 48
   RMM of potassium nitrate (KNO_3) = 101

   $$\text{Percentage of potassium (K)} = \frac{39}{101} \times 100\%$$
   101
   $$= \mathbf{38.6\%}$$

3. The compound is copper (II) sulphate ($CuSO_4.5H_2O$).
   Copper (Cu) = (1 x 64) = 64
   Sulphur (S) = (1 x 32) = 32
   Oxygen (O) = (4 x 16) = 64
   RMM of cupper(II) sulphate = 160

   Hydrogen (H) = (5 x 2) = 10
   Oxygen (O) = (5 x 16) = 80
   RMM of water(H_2O) = 90

   RMM of copper (II) sulphate ($CuSO_4.5H_2O$) = 160 + 90 = 250

   $$\text{Percentage of water}(5H_2O) = \frac{90}{250} \times 100\%$$
   $$= \mathbf{36\%}$$

4a. The compound is potassium nitrate (KNO_3).

Potassium (K) = (1 x 39) = 39
Nitrogen (N) = (1 x 14) = 14
Oxygen (O) = (3 x 16) = 48
RMM of potassium nitrate (KNO_3) = 101

$$\text{Percentage of nitrogen (N)} = \frac{14}{101} \times 100\%$$
$$= \mathbf{13.9\%}$$

b. The compound is ammonium chloride (NH_4Cl).

Nitrogen (N) = (1 x 14) = 14
Hydrogen (H) = (1 x 4) = 4
Chlorine (Cl) = (1 x 35.5) = 35.5
RMM of ammonium chloride (NH_4Cl) = 53.5

$$\text{Percentage of chlorine (Cl)} = \frac{35.5}{53.5} \times 100\%$$
$$= \mathbf{66.36\%}$$

Learning Activity 7

1. a. $$\text{number of moles of calcium (Ca)} = \frac{100}{40}$$
$$= \mathbf{2.5mol}$$

b. $$\text{number of moles of potassium (K)} = \frac{3.9}{39}$$
$$= \mathbf{0.1mol}$$

c. $$\text{number of moles of silicon (Si)} = \frac{70}{28}$$
$$= \mathbf{2.5mol}$$

d. $$\text{number of moles of flourine (Fl)} = \frac{9.5}{19}$$
$$= \mathbf{0.5mol}$$

2. a. $$\text{number of moles of sodium chloride (NaCl)} = \frac{19.5}{58.5}$$
$$= \mathbf{0.33mol}$$

b. $$\text{number of moles of aluminium chloride } (AlCl_3) = \frac{267}{133.5}$$

$= \textbf{2mol}$

c. number of moles of copper sulphate $(CuSO_4) = \frac{480}{160}$

$= \textbf{3mol}$

d. number of moles of silver nitrate $(AgNO_3) = \frac{42.5}{170}$

$= \textbf{0.25mol}$

Learning Activity 8

1. a. glucose solution $(C_6H_{12}O_6) = \frac{0.24mol}{0.50dm^3}$

$= \mathbf{0.48mol/dm^3}$

b. sodium chloride (NaCL) $= \frac{0.010mol}{0.20dm^3}$

$= \mathbf{0.05mol/dm^3}$

c. solution $= \frac{2x10^{-3}\,mol}{25.0dm^3}$

$= \mathbf{8 \times 10^{-5} mol/dm^3}$

d. solution $= \frac{1.23mol}{4.10dm^3}$

$= \mathbf{0.3mol/dm^3}$

2. a. ammonia (NH_3) = (3.4) (1.0)

= **3.4mol**

b. ethanol (C_2H_5OH) = (0.24) (0.45)

= **0.108mol**

c. ethanol (C_2H_5OH) = (6.0) (0.015)

= **0.09mol**

d. sucrose $(C_{12}H_{22}O_{11})$ = (2.5) (0.015)

= **0.0375mol**

Learning Activity 9

1. a. sulphur dioxide (SO_2) $= \dfrac{1.2}{24dm^3}$

$=$ **0.05mol**

b. methane (CH_4 $= \dfrac{0.24}{24dm^3}$

$=$ **0.01mol**

c. carbon dioxide (CO_2) $= \dfrac{0.12}{24dm^3}$

$=$ **5×10^{-3}mol**

d. sulphur dioxide (SO_2) $= \dfrac{0.5}{24dm^3}$

$=$ **0.021mol**

2. a. oxygen gas (O_2) $=$ (0.1) ($24dm^3$)

$=$ **$2.4dm^3$**

b. hydrogen (H_2) $=$ (3) ($24dm^3$)

$=$ **$72dm^3$**

c. chlorine (Cl_2) $=$ (5) ($24dm^3$)

$=$ **$120dm^3$**

d. sulphur dioxide (SO_2) $=$ (0.05) ($24dm^3$)

$=$ **$1.2dm^3$**

3. a. oxygen (O_2) $=$ (0.5) (22 400cm^3)

$=$ **11, 200cm^3**

b. hydrogen (H_2) $=$ (2) (22 400cm^3)

$=$ **44, 800cm^3**

c. chlorine (Cl_2) $=$ (3) (22 400cm^3)

$=$ **67, 200cm^3**

d. sulphur dioxide (SO_2) $=$ (0.25) (22 400cm^3)

$=$ **5 ,600cm^3**

4. a. sulphur dioxide (SO_2) $= \dfrac{1200}{22400cm^3}$

$= \mathbf{0.05mol}$

b. methane (CH_4) $= \dfrac{124}{22400cm^3}$

$= \mathbf{5.5 \times 10^{-3} mol}$

c. carbon dioxide (CO_2) $= \dfrac{1.50}{22.4dm^3}$

$= \mathbf{0.07mol}$

d. number of moles sulphur dioxide(SO_2) $= \dfrac{240}{22400cm^3}$

$= \mathbf{0.01mol}$

Learning Activity 10

1.

	C	:	H	:	Cl
Ratio of mass	2.4	:	0.2	:	21.3
Divide by atomic mass (A_r)	$\frac{2.4}{12}$	:	$\frac{0.2}{1}$	:	$\frac{21.3}{35.5}$
Divide by the smaller number	$\frac{0.2}{0.2}$	:	$\frac{0.2}{0.2}$	:	$\frac{0.6}{0.2}$
Ratio of atoms	1	:	1	:	3

The empirical formula of the sample is $\mathbf{CHCl_3}$.

2.

	N	:	O
Ratio of mass	1.58	:	3.62
Divide by atomic mass (A_r)	$\frac{1.58}{14}$	:	$\frac{3.62}{16}$
Divide by the smaller number	$\frac{0.11}{0.11}$	:	$\frac{0.23}{0.11}$

Ratio of atoms	1	:	2

The empirical formula of the oxide of nitrogen is **NO_2**.

3.

	C	:	H
Ratio of mass	93.7	:	6.3
Divide by atomic mass (A_r)	$\frac{93.7}{12}$	:	$\frac{6.3}{1}$
Divide by the smaller number	$\frac{7.8}{6.3}$	:	$\frac{6.3}{6.3}$
Ratio of atoms	1	:	1

The empirical formula of the naphthalene is **CH.**

4.

	Si	:	O
Ratio of mass	46.7	:	53.3
Divide by atomic mass (A_r)	$\frac{46.7}{28}$	:	$\frac{53.3}{16}$
Divide by the smaller number	$\frac{1.68}{1.68}$	:	$\frac{33.3}{1.68}$
Ratio of atoms	1	:	1.9 rounded to 2

The empirical formula of the compound is **SiO_2**.

5. The **empirical formula** of a compound can be found from its percentage composition. The empirical formula is the formula which only shows the simplest ratio of the elements present in a molecule.

Learning Activity 11

1a.

	C	:	H
Ratio of mass	7.2	:	1.2
Divide by atomic mass (A_r)	$\frac{7.2}{12}$	:	$\frac{1.2}{1}$
Divide by the smaller number	$\frac{0.6}{0.6}$	:	$\frac{1.2}{0.6}$
Ratio of atoms	1	:	2

The empirical formula is $\mathbf{CH_2}$.

b. The empirical formula mass is (1 x 12) + (2 x 1) = 14.
Since the molar mass of the compound is 84g/mol, divide it by the empirical formula mass. So, $\frac{84}{14} = 6$.
The molecular formula is $\mathbf{(CH_2)6 = C_6H_{12}}$.

2 . Since the empirical formula is CH_2 then its empirical formula mass is (1 x 12) + (1 x 2= 14).

The molar mass is 28g/mol, then 28/14 = 2. Hence, the molecular formula is $(CH_2)_2$ = $\mathbf{C_2H_4}$

3.

	C	:	H	:	O
Ratio of mass	40%	:	6.6%	:	53.3%
Divide by atomic mass (A_r)	$\frac{40}{12}$	:	$\frac{6.6}{1}$	:	$\frac{53.3}{16}$
Divide by the smaller number	$\frac{3.3}{3.3}$	:	$\frac{6.6}{3.3}$	:	$\frac{3.3}{3.3}$
Ratio of atoms	1	:	2	:	1

The empirical formula of nicotine is $\mathbf{CH_2O}$.
Since the empirical formula is CH_2O then its empirical formula mass is (1x 12) + (1x 2) + (1 x 16) = 30.
The molar mass is 180 g/mol, then 180/30 = 6. Hence, the molecular formula is 6 x CH_2O = $\mathbf{C_6H_{12}O_6}$.

4. C : H : O : N

Ratio of mass	49.1 :	5.1 :	16.5 :	28.9
Divide by atomic mass (A_r)	$\frac{49.1}{12}$:	$\frac{5.1}{1}$:	$\frac{16.5}{16}$:	$\frac{28.9}{14}$
Divide by the smaller number	$\frac{4.09}{1.03}$:	$\frac{5.1}{1.03}$:	$\frac{1.03}{1.03}$:	$\frac{2.06}{1.03}$
Ratio of atoms	4 :	5 :	1	2

The empirical formula of nicotine is $\mathbf{C_4H_5ON_2}$.

Since the empirical formula is $C_4H_5ON_2$ then its empirical formula mass is (4 x 12) + (1 x 5) + (1 x 16) + (2 x 14) = 97.

The molar mass is 195 g/mol, then 195/97 = 2. Hence, the molecular formula is 2 x $C_4H_5ON_2$ = $\mathbf{C_8H_{10}O_2N_4}$.

5. The **molecular formula** is the true formula which shows all the atoms present in a molecule. It can be calculated from the empirical formula and the relative molecular mass.

Learning Activity 12

1. a. $n(Mg) = \frac{6}{24}$

$= 0.25mol$

mass (Mg) = number of moles x molar mass

= 0.25 x 40 (MgO = (1 x 24) + 16 = 40)

= **10g**

b. From equation:

2 moles of magnesium (Mg) reacts with 1 mole of oxygen (O_2).

0.25 moles of magnesium (Mg) reacts with $\frac{0.25}{2}$ = 0.12mol of oxygen (O_2).

- Volume of oxygen at **r.t.p.** = number of moles x $24dm^3$
  = 0.12 x 24
  = $\mathbf{2.88dm^3}$
- Volume of oxygen **at s.t.p.** = number of moles x $22.4dm^3$

$= 0.12 \times 22.4dm^3$
$= \mathbf{2.68dm^3}$

2. a. $n(NaCl) = \frac{5.85}{58.5mol}$

$= 0.1mol$

$m(AgNO_3) = n \times Mr\ (AgNO_3)$
$= 0.1 \times 170$ — $M_r\ AgNO_3$ is $(1 \times 108) + (1 \times 14) + (3 \times 16) = 170$.
$=$ **170g**

b. $m\ (AgCl) = n \times M_r\ (AgCl)$
$= 0.1 \times 143.5$
M_r of AgCl $(1 \times 108) + (1 \times 35.5) = 143.5$.
$=$ **14.35g**

3. The number of moles of copper hydroxide ($Cu(OH)_2$) is:

$n(Cu(OH)_2) = \frac{10}{98}$
$= 0.10mol$

The mass of copper chloride ($CuCl_2$) is:

$m(CuCl_2) = n \times M_r$
$= 0.10 \times 135$
$=$ **13.50g**

The mass of sodium hydroxide is:

$n(NaOH) = 2 \times n(Cu(OH)_2)$
$= 2 \times 0.10$
$= 0.20\ mol$

$m(NaOH) = n \times M_r$
$= 0.20 \times 40$
$=$ **8.0g**

4. a. $n(CuO) = \frac{m}{M_r}$
$= \frac{5.0}{80}$
$=$ **0.0625mol**

The mass of copper (Cu) is:

m (Cu) = n x M_r
= 0.0625 x 64
= **4g**

b. Volume of hydrogen gas (H_2) at s.t.p = n x $22.4dm^3$
= 0.0625 x 22.4
= $\mathbf{1.4dm^3}$

5. $n = \frac{m}{M_r}$

$= \frac{50.0}{233}$ Mr of $BaSO_4$ = (1 x 137) + (1 x 32) + (4 x 16) = 233)

= **0.21mol**

Learning Activity 13

1. a. Number of moles of sodium chloride (NaCl) $= \frac{20}{58.5}$

= **0.342 mol**

Number of moles of silver nitrate ($AgNO_3$) $= \frac{20}{170}$

= **0.118 mol**

The sodium chloride (NaCl) is in excess. Silver nitrate ($AgNO_3$) is the limiting reactant.

b. Mass of silver chloride (AgCl) = (0.118) (143.5)
= **16.93g**

c. Mass of sodium chloride(NaCl) = (0.118) (58.5)
= 6.903g

Hence, **20g-6.903g = 13.097g of sodium chloride (excess reactant).**

2. a. Number of moles of zinc (Zn) $= \frac{20}{65}$

= **0.3mol**

Number of moles of copper sulphate (CuSO4) $= \frac{64}{160}$
= **0.4mol**

Copper sulphate ($CuSO_4$) is the excess reactant. Zinc (Zn) is the limiting reactant.

b. Mass of copper (Cu) = (0.3) (64)
= **19.2g**

c. Mass of zinc sulphate ($ZnSO_4$) = (0.3) (161)
= **48.3g**

d. mass of excess copper sulphate($CuSO_4$) = (0.1) (160)
= **16g**

Learning Activity 14

1. a. Number of moles of copper(Cu) $= \frac{6.4}{64}$ (A_r of copper)
$=$ 0.1 mol

Mass of copper oxide (CuO) = (0.1) (80) Mr of CuO is 80.
= **8 g copper oxide**

b. Percentage yield $= \frac{7.6}{8} \times 100\%$
= **95%**

2a. Number of mole of nitrogen (N_2) $= \frac{28}{28}$ M_r of N_2 is 28
= 1 mol

Mass of ammonia (NH_3) = (1) (17) Mr of NH_3 is 80
= **17g**

b. Percentage yield of ammonia (NH_3) $= \frac{3.4}{17} \times 100\%$
= **20%**

3a. Number of moles of sulphuric acid(H_2SO_4) $= \frac{98}{98}$
= **1mol**

Theoretical yield of ammonium sulphate $(NH_4)_2SO_4$ = (1) (132)
= **132g**

b. % yield of ammonium sulphate $(NH_4)_2SO_4$ $= \frac{120}{132} \times 100\%$
= **91%**

4. a. nitrogen + hydrogen $\longrightarrow$ ammonia

$$N_{2(g)} + 3H_{2(g)} \longrightarrow 2NH_{3(g)}$$

b. Number of moles of nitrogen (N_2) $= \frac{28}{28}$

$=$ **1mol**

Number of moles of hydrogen (H_2) $= \frac{8}{2}$

$=$ **4mol**

Hydrogen (H_2) is the **excess reactant**. **Nitrogen (N_2)** is the **limiting reactant.**

c. Theoretical yield of ammonia (NH_3) $=$ (1) (17)

$=$ **17g**

Percentage yield of ammonia (NH_3) $= \frac{5.1}{17} \times 100\%$

$=$ **30%**

Learning Activity 15

a. $NaOH_{(aq)} + HCl_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$

b. **0.025 dm^3**

c. n $=$ c x V

$=$ 0.100 x 0.025

$=$ **2.5 x 10^{-3}mol**

Learning Activity 16

1. $C_2 = \frac{C_1V_1}{V_2}$

$= \frac{(0.20)(0.020)}{0.250}$

$=$ **0.016 mol/dm^3 or 0.016M**

2. $V_2 = \frac{C_1V_1}{C_2}$

$$= \frac{(12.5)(0.01)}{0.2}$$

$$= \mathbf{0.625dm^3}$$

3. Before dilution:

$C_1V_1 = C_2V_2$

$$C_1 = \frac{C_1V_1}{V_1}$$

$$= \frac{(0.1)(1)}{0.25}$$

$$= \mathbf{0.4mol/L}$$

After dilution:

$C_1V_1 = C_2V_2$

$$C_2 = \frac{C_1V_1}{V_2}$$

$$= \frac{(0.4)(0.25)}{0.1}$$

$$= \mathbf{0.4mol/L}$$

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