Ahmad Ibrahim Secondary School

Name:_________________________________ (   ) Class:________ Date:__________

Chapter 9 : The Mole

Formula, Stoichiometry, Mole Concept
- Relative Atomic Mass
- Relative Molecular/ Formula Mass
- Percentage Composition of an Element in a Compound
- Mass of an Element in a given mass of a Compound

A Relative Atomic Mass (A_r)

At the end of this section, you should be able to:
- *Define relative atomic mass, A_r.*

Definition:

The relative atomic mass (A_r) of an atom is the **average** mass of one atom of that element compared to $\frac{1}{12}$ **of the mass of one carbon-12 atom**.

For example:

Chlorine consists of 2 isotopes → Cl-35 (75%) and Cl-37 (25%)

A_r of each chlorine atom = $\frac{(75 \times 35)+(25 \times 37)}{100} = 35.5$

Check Point 1

Calculate the A_r of the following elements, with the given isotopic abundances.

	Element	Isotopes	Relative Abundance	A_r of Element
(a)	Carbon	C-12	98.89%	$\frac{(98.9 \times 12)+(1.11 \times 13)}{100}$ = 12.0111 = 12.0
		C-13	1.11%	
(b)	Neon	Ne-20	90.5%	$\frac{(90.5 \times 20)+(0.3 \times 21)+(9.2 \times 22)}{100}$ = 20.187 = 20.2
		Ne-21	0.3%	
		Ne-22	9.2%	
(c)	Potassium	K-39	93.38%	$\frac{(93.38 \times 39)+(0.01 \times 40)+(6.61 \times 41)}{100}$ = 39.1323 = 39.1
		K-40	0.01%	
		K-41	6.61%	

From the Periodic Table,

Atomic (mass) number = Relative atomic mass

$$^{12}_{6}C$$

Proton number

Check Point 2

Write down the relative atomic mass for the following elements:

(a)	Sodium, Na	23	(b)	Calcium, Ca	40
(c)	Oxygen, O	16	(d)	Copper, Cu	64

B Relative Molecular/ Formula Mass (M_r)

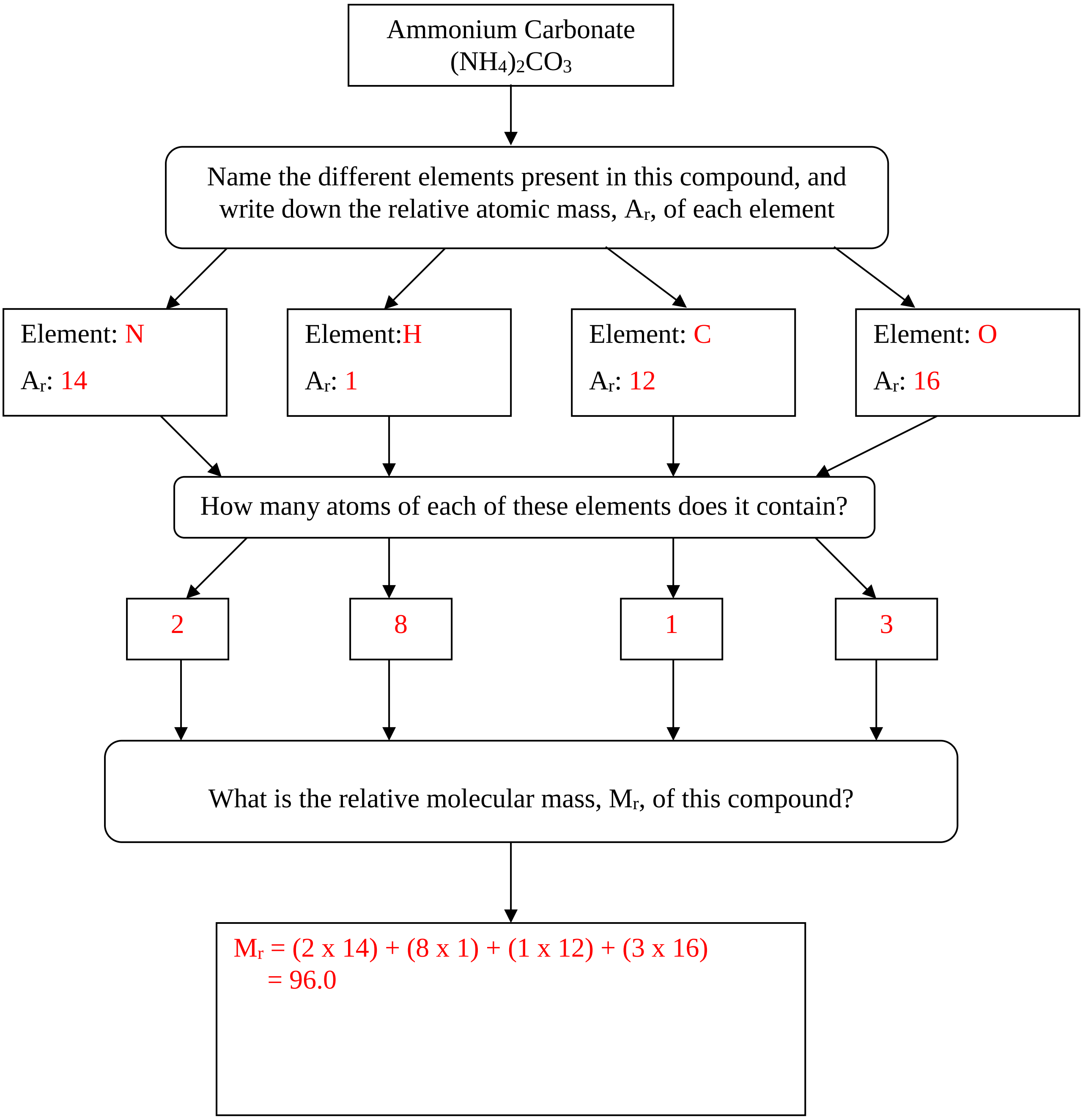
At the end of this section, you should be able to:-

- *Define relative molecular mass, M_r*
- *Calculate relative molecular mass (and relative formula mass) as the sum of relative atomic masses*

Definition:

The relative molecular / formula mass of a compound is the **average** mass of one molecule of that element of compound compared to $\frac{1}{12}$ **of the mass of one carbon-12 atom**.

For example:


Ammonium Carbonate
$(NH_4)_2CO_3$
Name the different elements present in this compound, and write down the relative atomic mass, A_r, of each element
Element: N
A_r: 14
Element:H
A_r: 1
Element: C
A_r: 12
Element: O
A_r: 16
How many atoms of each of these elements does it contain?
2
8
1
3
What is the relative molecular mass, M_r, of this compound?
M_r = (2 x 14) + (8 x 1) + (1 x 12) + (3 x 16)
= 96.0

Check Point 3

1 Find the relative molecular mass of each of the following compounds:

(a) Carbon dioxide, CO_2

$12 + 16 + 16 = 44.0$

(b) Ammonium chloride, NH_4Cl

$14 + 4 + 35.5 = 53.5$

(c) Ammonium phosphate, $(NH_4)_3PO_4$

$(18 \times 3) + 31 + 4(16) = 149$

(d) $MgSO_4.7H_2O$

$24 + 32 + 4(16) + 7(18) = 246$

2 Calculate the value of x in $CuSO_4.xH_2O$, given the M_r of $CuSO_4.xH_2O$ is 250.

Mr of $CuSO_4.xH_2O$ = 250
Mr of xH_2O = 250 – 160 = 90
$x = \frac{90}{18} = 5$

C Percentage Composition of an Element in a Compound

At the end of this section, you should be able to:-

- *Calculate % composition of an element in a compound when given appropriate information*

To calculate the percentage composition of an element in a compound:

Percentage composition of an element in a compound

$$= \frac{\textit{number of atoms of element x Ar}}{\textit{Mr of compound}} \times 100\%$$

<u>Example 1:</u>

Find the percentage composition of copper in copper(II) sulfate $CuSO_4$.

$$\text{percentage composition of copper} = \frac{1 \, x \, 64}{64 + 32 + 4(16)} \, x \, 100\% = 40.0\%$$

<u>Example 2:</u>

Find the percentage composition of hydrogen in ethanoic acid, CH_3COOH.

percentage composition of hydrogen $= \dfrac{1 \, x \, 4}{12+3+12+16+16+1} x \; 100\% = 6.67\%$

<u>Example 3:</u>

Hydrated copper(II) sulfate crystals ($CuSO_4.5H_2O$) can be dehydrated by heating. Calculate the percentage composition by mass of water present in the molecule.

percentage composition of water $= \dfrac{5 \, x \, 18}{160+5(18)} x \; 100\% = 36.0\%$

D Mass of an Element in given mass of a Compound

At the end of this section, you should be able to:-

- *Calculate mass of an element in a compound when given the mass of the compound*

To calculate the mass of an element in a given mass of a compound:

Mass of an element in a given mass of a compound

$$= \frac{\textbf{number of atoms of element } x \textbf{ Ar}}{\textbf{Mr of compound}} \times \textbf{mass of compound}$$

<u>Example 1:</u>

Calculate the mass of Cu in 32 g of $CuSO_4$.

Mass of Cu $= \frac{64}{160} x \; 32 = 12.8g$

<u>Example 2:</u>

Calculate the mass of water in 12.3 g of $MgSO_4.7H_2O$.

Mass of water $= \frac{7 \, x \, 18}{24+32+4(16)+(7x18)} x 12.3 = 6.3$

Formula, Stoichiometry, Mole Concept
Worksheet 9A

Answer the following questions.

1 Calculate the M_r of the following substances.

(a) Na_2O 23 + 23 + 16 = 62.0

(b) H_2SO_4 2+ 32+ 4(16) = 98.0

2 Calculate the percentage composition of the following elements in the respective compounds.

(a) Fe in $FeCO_3$

$$\frac{56}{116} x\ 100\% = 48.3\%$$

(b) O in $NaNO_3$

$$\frac{3(16)}{85} x\ 100\% = 56.5\%$$

(c) H in $MgSO_4.7H_2O$

$$\frac{14}{246} x\ 100\% = 5.69\%$$

3 A magnesium strip produces magnesium oxide when burnt in air.

(a) Write the chemical formula of magnesium oxide. MgO

(b) Calculate the percentage composition of magnesium in magnesium oxide.

$$\frac{24}{40} x\ 100\% = 60.0\%$$

Formula, Stoichiometry, Mole Concept
- The Mole
- Calculate Empirical Formula of a Compound
- Calculation of Molecular Formula
- Molar mass, and its relationship with moles and mass
- Molar volume, and its relationship with moles and volume

E The Mole https://www.youtube.com/watch?v=TEl4jeETVmg

At the end of this section, you should be able to:-
- *Measure the number of atoms, molecules or ions in moles*
- *Use the Avogadro's constant to calculate the number of moles and number of particles*

- When we say:
  - A pair of oranges => 2 oranges
  - A dozen of eggs => 12 eggs

- In Chemistry, when we say one mole = 6×10^{23} particles (This is also called the **Avogadro's constant**.)
  - Particles can be atoms, molecules, compounds etc.
  - For example:

    1 mole of carbon dioxide gas = 6×10^{23} carbon dioxide molecules

    0.5 mole of carbon dioxide gas = 3×10^{23} carbon dioxide molecules

    Likewise,

    1 mol of hydrogen gas = 6×10^{23} hydrogen molecules

    2 mol of hydrogen gas = 1.2×10^{24} hydrogen molecules

- What does it mean in a chemical equation?

  Example: $2\ H_2 + O_2 \rightarrow 2\ H_2O$

Instead of stating that 12 x 10^{23} hydrogen molecules react with 6 x 10^{23} oxygen molecules to produce 12 x 10^{23} water molecules, we can just simply state that 2 moles of hydrogen molecules react with 1 mole of oxygen molecule to produce 2 moles of water molecules.

F The Molar Mass

At the end of this section, you should be able to:-
- *Define molar mass*

The molar mass of an element / *molecule* is the mass of **one mole of atoms** of the element / *molecule*. It is **equal to its relative atomic mass (A_r)** / **relative molecular mass (Mr)**.

G Calculation of MOLE

$$\boxed{number\ of\ moles = \frac{mass\ (g)}{molar\ mass\ (\frac{g}{mol})}}$$

Ar or Mr
from Periodic Table

1 Calculate the number of moles in

(a) 64 g of oxygen gas O_2

Number of mole of oxygen = $\frac{64}{16\ x\ 2} = 2.00$

(b) 12.25 g of barium nitrate $Ba(NO_3)_2$

Number of mole of barium nitrate = $\frac{12.25}{261} = 0.0469$

Mass = Moles x Ar or Mr

2 Calculate the mass of

(a) 3 moles of chlorine atoms

Mass of chlorine atoms = 3 x 35.5 = 106.5 g

(b) 0.2 mole of bromine gas

Mass of bromine gas = 0.2 x (80 x 2) = 32 g

(c) 0.6 moles of sulfuric acid

Mass of sulfuric acid = 0.6 x 98 = 58.8 g

H Calculate Empirical Formula of a Compound

At the end of this section, you should be able to:-
- *Calculate empirical formulae from relevant data*

What is Empirical Formula?

Sugar (glucose) has a molecular formula of $C_6H_{12}O_6$. simplest form

Its empirical formula is CH_2O, which shows the **ratio** of the elements in the compound.

Example 1:

A compound consists of 7.0 g of nitrogen combined with 4.0 g of oxygen. Find the formula of the compound.

	N	O
Mass	7	4
No. of moles $\frac{mass\ (g)}{molar\ mass\ (\frac{g}{mol})}$	$\frac{7}{14} = 0.5$	$\frac{4}{16} = 0.25$
Simplest Ratio (divide by smallest mole number)	$\frac{0.5}{0.25} = 2$	$\frac{0.25}{0.25} = 1$

Therefore, empirical formula: N_2O

Example 2:

12.0 g of anhydrous magnesium sulfate combines with 12.6 g of water to form hydrated magnesium sulfate. What is the formula of the hydrated magnesium sulfate? $MgSO_4.7H_2O$

	$MgSO_4$	H_2O
Mass	12	12.6
No. of moles $\frac{mass\ (g)}{molar\ mass\ (\frac{g}{mol})}$	$\frac{12}{120} = 0.1$	$\frac{12.6}{18} = 0.7$
Simplest Ratio (divide by smallest mole number)	$\frac{0.1}{0.1} = 1$	$\frac{0.7}{0.1} = 7$

Check Point 4

1 When a 5.76 g of a compound was analyzed, it was found to be an oxide of copper containing 5.12 g of copper. Calculate the empirical formula of this oxide. **Cu_2O**

	Cu	O
Mass	5.12	5.76 – 5.12 = 0.64
No. of moles $\frac{mass\ (g)}{molar\ mass\ (\frac{g}{mol})}$	$\frac{5.12}{64} = 0.08$	$\frac{0.64}{16} = 0.04$
Simplest Ratio (divide by smallest mole number)	$\frac{0.08}{0.04} = 2$	$\frac{0.04}{0.04} = 1$

2 A compound has the composition by mass: 29.4% calcium, 23.5% sulfur and 47.1% oxygen. Find the empirical formula.

	Ca	S	O
Mass	29.4	23.5	47.1
No. of moles	$\frac{29.4}{40} = 0.735$	$\frac{23.5}{32} = 0.734$	$\frac{47.1}{16} = 2.94$
Simplest Ratio (divide by smallest no.)	$\frac{0.735}{0.734} = 1$	$\frac{0.734}{0.734} = 1$	$\frac{2.94}{0.734} = 4$

Therefore, empirical formula: **$CaSO_4$**

3 Barium chloride, $BaCl_2$ forms a hydrate which contains 85.25% barium chloride and 14.75% water of crystallization. What is the formula of hydrated barium chloride?

	$BaCl_2$	H_2O
Mass	85.25	14.75
No. of moles	$\frac{85.25}{208} = 0.410$	$\frac{14.75}{18} = 0.819$
Simplest Ratio (divide by smallest mole number)	$\frac{0.41}{0.41} = 1$	$\frac{0.819}{0.41} = 2$

$BaCl_2.\underbrace{2H_2O}_{\text{water of crystallisation}}$ hydrated salt

powder vs crystals

anhydrous hydrated

I Calculation of Molecular Formula

At the end of this section, you should be able to:-
- *Calculate molecular formulae from relevant data*

Empirical formula shows the **simplest ratio** between the atoms in a compound.

Molecular formula shows the **exact number of atoms** of each element in a molecule.
chemical

Substance	Molecular formula	Relative molecular mass	Empirical formula	Relative mass (from empirical formula)	$\frac{Mr\ (molecular\ formula)}{Mr\ (empirical\ formula)}$
Water	H_2O	18.0	H_2O	18	1
Ammonia	NH_3	17.0	NH_3	17	1
Hydrogen peroxide	H_2O_2	34.0	HO	17	2
Ethane	C_2H_6	30.0	CH_3	15	2
Ethene	C_2H_4	28.0	CH_2	14	2
Propene	C_3H_6	42.0	CH_2	14	3

If empirical formula is A_xB_y,

Molecular formula is $(A_xB_y)_n$ where $_n$ = 1,2,3….

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

Example 1:

The empirical formula of a compound is COH_3. Its relative molecular mass is 62. Find the molecular formula.

Mr of COH_3 = 12 + 16 + 3 = 31.0

Let the molecular formula be x(COH_3).

Mr of x(COH_3) = 62

x(31) = 62

x = 2

molecular formula is **$C_2O_2H_6$**

	empirical formula	molecular formula
	COH_3	? $C_2O_2H_6$
Mr	12+16+3(1) = 31	62

×2

Check Point 5

1 Caffeine is a compound found in coffee and tea. The percentage composition of caffeine is 49.5% carbon, 5.1% hydrogen, 15.6% oxygen and 28.9% nitrogen. The relative molecular mass of caffeine is 195. Determine

(a) the empirical formula and

(b) the molecular formula of caffeine.

	C	H	O	N
Mass	49.5	5.1	15.6	28.9
No. of moles	$\frac{49.5}{12} = 4.125$	$\frac{5.1}{5.1} = 5.1$	$\frac{15.6}{16} = 0.975$	$\frac{28.9}{14} = 2.06$
Simplest Ratio	$\frac{4.125}{0.975} = 4.2$	$\frac{5.1}{0.975} = 5.23$	$\frac{0.975}{0.975} = 1$	$\frac{2.06}{0.975} = 2.11$

(a) empirical formula = $C_4H_5ON_2$

(b) Let the molecular formula be x($C_4H_5ON_2$).

Mr of x($C_4H_5ON_2$) = 195

x(97) = 195

x = 2

molecular formula is $C_8H_{10}O_2N_4$

J Molar Volume of GASES

At the end of this section, you should be able to:-

- *Use molar volume to calculate the number of moles and the volume of **gases** in questions involving gases/ gas reaction*

O-levels

- Molar volume is the volume occupied by **one mole of gas** **at room temperature and pressure**.

➔ **24** dm^3

$$\boxed{\text{number of moles of } \textbf{gas} = \frac{\text{volume of gas in } dm^3}{24 dm^3}}$$

Note that $1 dm^3 = 1000 cm^3$

Example 1:

(a) Find the number of moles in 12 dm^3 of oxygen gas.

Number of moles of oxygen gas = 12/24 = 0.500

$\frac{12}{24}$ actual vol. molar volume

(b) Find the volume of 0.8 moles of carbon dioxide gas.

Volume of CO_2 = 0.8 x 24 = 19.2 dm^3

mo

Check Point 6

1 What is the volume (at r.t.p.) would the following gas occupy?

(a) 0.75 mol of hydrogen gas.

Volume = 0.75 x 24 = 18.0 dm^3

(b) 1.5 mol of methane.

Volume = 1.5 x 24 = 36.0 dm^3

2 How many moles does each of the following gases, at r.t.p., contain?

(a) 3.6 dm^3 of nitrogen gas.

Number of moles = 3.6/24 = 0.150 mol

(b) 2400 cm^3 of sulfur dioxide gas (remember to change to dm^3)

Number of moles = 2.4/24 = 0.100

3 Calculate the volume (at r.t.p.) occupied by 34 g of ammonia gas.

Number of mole of ammonia gas = 34 / 17 = 2

Volume = 2 x 24 = 48.0 dm^3

Worksheet 9B

Answer the following questions on writing paper.

1 Find the empirical formula of a compound which contains 40% sulfur and 60% oxygen by mass.

2 (a) Calculate the percentage composition by mass of tin in tin(II) chloride ($SnCl_2$).

(b) Tin reacts with chlorine to form another chloride containing 45.6% of tin.

(i) What is the empirical formula of this chloride?

(ii) If the M_r of this chloride of tin is 261. What is the molecular formula of the chloride?

3 The following results were obtained in an experiment to determine the formula of an oxide of silicon.

Mass of crucible	= 15.20 g
Mass of crucible + silicon	= 15.48 g
Mass of crucible + oxide of silicon	= 15.80 g

(a) Find the empirical formula of the oxide of silicon.

(b) If the M_r of the oxide of silicon is 60, what is its molecular formula?

4 A sample of hydrated copper(II) sulfate weighs 124.8 g. the sample has been determined to contain 31.8 g of copper(II) ions and 48.0 g of sulfate ions.

(a) How many molecules of water of crystallization are present in the sample?

(b) Deduce the actual formula of hydrated copper(II) sulfate.

Answer

(1)

	S	O
Mass	**40**	**60**
No. of moles	$\frac{40}{32} = 1.25$	$\frac{60}{16} = 3.75$
Simplest Ratio (divide by smallest mole number)	$\frac{1.25}{1.25} = 1$	$\frac{3.75}{1.25} = 3$

Empirical formula = SO_3

(2a) % mass of Sn = 119 / 190 x 100% = 62.6%

(2b)

	Sn	Cl
Mass	**45.6**	**54.4**
No. of moles	$\frac{45.6}{119} = 0.383$	$\frac{54.4}{35.5} = 1.53$
Simplest Ratio	$\frac{0.383}{0.383} = 1$	$\frac{1.53}{0.383} = 4$

Empirical formula = $SnCl_4$

(2bii)

Let the molecular formula be x($SnCl_4$).

Mr of x($SnCl_4$) = 261

x(261) = 261

x = 1

Molecular formula = $SnCl_4$

(3a)

	Si	**O**
Mass	**0.28**	**0.32**
No. of moles	$\frac{0.28}{28} = 0.01$	$\frac{0.32}{16} = 0.02$
Simplest Ratio	**1**	**2**

Empirical formula = SiO_2

(3b)

Let the molecular formula be x(SiO_2).

Mr of x(SiO_2) = 60

x(60) = 60

x = 1

Molecular formula = SiO_2

(4a)

Mass of water of crystallization molecules = 124.8 – 31.8 – 48 = 45.0g

Number of mole of water of crystallization molecules = 45 / 18 = 2.5 mol

Number of water of crystallization molecules = 2.5 x 6 x 10^{23} =**<u>1.5 x 10^{24}</u>**

(4b)

	Cu	**SO_4**	**H_2O**
Mass	**31.8**	**48**	**45**
No. of moles	$\frac{31.8}{64} = 0.497$	$\frac{48}{96} = 0.5$	$\frac{45}{18} = 2.5$
Simplest Ratio	**1**	**1**	**5**

Formula: $CuSO_4.5H_2O$

K Concentration of Solutions

At the end of this section, you should be able to:-
- *Calculate the concentration of solutions*
- *Apply the concept of solution concentration (mol/ dm^3 or g/ dm^3) to solve simple problems*

Concentration

- Concentration of a solution is given by **the amount of a solute (in g or mol) dissolved in a unit volume (dm^3) of a solvent.**
- It can be expressed in **g/dm^3** or **mol/dm^3**
  - For example:
    - 50 g/dm^3 of sodium chloride solution means 50 g of NaCl in 1 dm^3 of water.
      Therefore, in 4 dm^3, there are 200 g of sodium chloride
    - 3 mol/dm^3 of sodium chloride solution means 3 mol of NaCl in 1 dm^3 of water.
      Therefore, in 5 dm^3, there are 15 mol of sodium chloride

Concentration of a Solution

Concentration in g/dm^3 = $\frac{mass\ of\ solute\ (g)}{volume\ of\ solvent\ (dm^3)}$

Concentration in mol/dm^3 = $\frac{number\ of\ moles\ of\ solute\ (mol)}{volume\ of\ solvent\ (dm^3)}$

Example 1:

Calculate the concentration of hydrochloric acid solution in mol/dm^3 if 400 cm^3 of the solution contains 0.25 mole of HCl.

Concentration = $\frac{0.25}{400/1000} = 0.625\ mol/dm^3$

Example 2:

Calculate the number of moles of sodium hydroxide contained in 150 cm^3 of 2.0 mol/dm^3 solution of sodium hydroxide.

Number of moles = $\frac{150}{1000} x\ 2 = 0.300\ mol$

Example 3:

A solution of sodium hydroxide contains 9.6 g of sodium hydroxide in 100 cm^3 of solution.
Find the concentration of the solution in

(a) g/dm^3;
(b) mol/dm^3

(a) Concentration in $g/dm^3 = 9.6 / \frac{1000}{100} = 96\ g/dm^3$

(b) Concentration in $mol/dm^3 = \frac{96}{40} = 2.4\ mol/dm^3$

Good for teachers to summarize the different formulas for finding mole before continuing to teach.

L Calculation from Equations

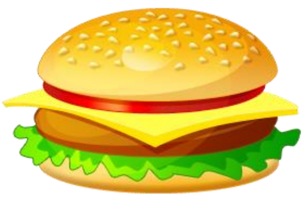
At the end of this section, you should be able to:-

- *Calculate stoichiometric reacting masses and volumes of gases (one mole of gas occupies 24dm^3 at room temperature and pressure)*

Let's consider this recipe:

CHICKEN & HAM BURGER

INGREDIENTS:
1 1 burger bun
2 2 slice of ham
3 1 slice of chicken patty
4 3 slices of tomato
5 4 pieces of lettuce
6 Honey mustard or other topping



This recipe is used to make 1 chicken and ham burger, how many slices of ham and chicken patty must we prepare to make 5 burgers?

Slices of ham required:**2 x 5 =10**

Slices of chicken patty required:**1 x 5 = 5**

A chemical equation is like a recipe, which tells us

- What substances react together (the **reactants**)
- What substances are produced in the reaction (the **products**)
- How much of each substances required

For example:

From this chemical equation,

$$2H_2 + O_2 \rightarrow 2H_2O$$

What does it tell you?

I need **2 mol** of hydrogen gas and **1 mol** of oxygen gas to produce **2 mol** of water

So if

- I have 4 moles of hydrogen and 2 moles of oxygen, I would get **4** mol of water.
- I have 3 moles of oxygen, I would get **6** mol of water.
- I want 0.5 moles of water, I would need **0.25** mol of oxygen.

Example 1:

Calculate the mass of water produced when 0.5 g of hydrogen gas is burnt in oxygen.

$$2H_2 + O_2 \rightarrow 2H_2O$$

Amount of hydrogen gas = $\frac{0.5}{2} = 0.250\ mol$

$2H_2 \equiv 2H_2O$ (ratio is 1:1)

Amount of water = 0.25 mol

Mass of water = 0.25 x 18 = 4.50g

Example 2:

Magnesium reacts with hydrogen chloride according to the equation:

$$Mg + 2HCl \rightarrow MgCl_2 + H_2$$

Calculate the mass of hydrogen chloride, HCl required to make 20 g of magnesium chloride, $MgCl_2$.

Amount of magnesium chloride = $\frac{20}{95} = 0.2105\ mol$

$2HCl \equiv MgCl_2$ (ratio is 2:1)

Amount of HCl = 0.2105 x 2 = 0.421 mol

Mass of HCl = 0.421 x 36.5 = 15.4 g

Example 3:

Nitrogen gas and hydrogen gas reacts together to form ammonia gas.

$$1\ N_2 + 3\ H_2 \rightarrow 2\ NH_3$$

Balance the equation above and calculate the volume of ammonia gas produced when 12 dm^3 of nitrogen gas is used.

Volume ratio ≡ mole ratio

$N_2 \equiv 2NH_3$

Volume of ammonia = 12 x 2 = 24dm^3

Check Point 7

1 Sulfuric acid and sodium hydroxide can react as follows:

$$H_2SO_4 + 2NaOH \rightarrow Na_2SO_4 + 2H_2O$$

(a) Calculate the relative formula mass of sodium sulfate and sodium hydroxide.

(b) What is the mass of sodium sulfate that can be formed from 40 g of sodium hydroxide?

(a) Na_2SO_4 = 142 ; NaOH = 40.0

(b) Amount of NaOH = 40/40 = 1 mol

$2NaOH \equiv Na_2SO_4$

Amount of Na_2SO_4 = 0.5 mol

Mass of Na_2SO_4 = 0.5 x 142 = 71 g

2 Magnesium ribbon burns in excess air to form magnesium oxide.

$$2Mg\ (s) + O_2\ (g) \rightarrow 2MgO\ (s)$$

Calculate the mass of magnesium oxide that would be produced if 60 g of magnesium were used.

Amount of Mg = 60/24 = 2.5 mol

$2Mg \equiv 2MgO$

Amount of MgO = 2.5 mol

Mass of MgO = 2.5 x (24+16) = 100 g

3 63.0 dm^3 of carbon monoxide, measured at r.t.p., was used to react with iron(III) oxide. What mass of iron was produced at the end of the reaction?

$$Fe_2O_3 (s) + 3 CO (g) \rightarrow 2 Fe (s) + 3 CO_2 (g)$$

Amount of CO = 63/24 = 2.625 mol

3CO ≡ 2Fe

Amount of Fe = 2.625 x $\frac{2}{3}$ = 1.75 mol

Mass of Fe = 1.75 x (56) = 98.0 g

Formulae, Stoichiometry, Mole Concept
Worksheet 9C

1. A solution contains 4.9 g of H_2SO_4 in 1.0 dm^3. Calculate the concentration in mol/dm^3.
   Amount of H_2SO_4 = 4.9/98 = 0.05 mol
   Concentration of H_2SO_4 = 0.05/1 = 0.0500 mol/dm^3

2. Calculate the number of moles of $FeSO_4$ in 200 cm^3 of 0.10 mol/dm^3 solution.
   Number of moles of $FeSO_4$ = 200/1000 x 0.1 = 0.0200

3. The formula of naphthalene is $C_{10}H_8$. It burns in air according to the equation below:

   $$C_{10}H_8 (s) + 12O_2 (g) \rightarrow 10CO_2 (g) + 4H_2O (l)$$

   16 g of naphthalene was burnt in excess oxygen. Calculate the mass of water produced in the combustion.

   Amount of naphthalene = 16/128 = 0.125 mol

   $C_{10}H_8 \equiv 4H_2O$

   Amount of water = 0.125 x 4 = 0.5 mol

   Mass of water = 0.5 x 18 = 9.00 g

4. Iron can be extracted from its ore, Fe_2O_3, by heating its ore with carbon strongly.

   $$2Fe_2O_3 + 3C \rightarrow 4Fe + 3CO_2$$

   Calculate the mass of pure iron that can be obtained from 40 g of iron ore?

   Amount of Fe_2O_3 = 40/160 = 0.25

   $2Fe_2O_3 \equiv 4Fe$

   Amount of Fe = 0.25 x 2 = 0.50 mol

   Mass of Fe = 0.5 x 56 = 28.0g

5. 14 g of nitrogen gas, N_2, was reacted with hydrogen, H_2, to produce ammonia, NH_3.

   (a) Write the balanced equation for the reaction.

   (b) What volume of ammonia is produced in the reaction?

(a) $N_2 + 3H_2 \rightarrow 2NH_3$

(b) Amount of nitrogen gas = 14/28 = 0.5 mol

$N_2 \equiv 2NH_3$

Amount of ammonia = 0.5 x 2 = 1 mol

Volume of ammonia = 1 x 24 = 24dm^3

Formulae, Stoichiometry, Mole Concept

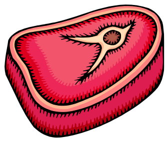
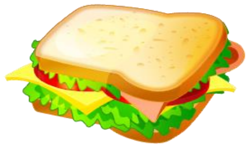
- Limiting and Excess Reactants
- Percentage Yield and Purity
- Titration Calculations

M Limiting and Excess Reactants

At the end of this section, you should be able to:-

- *Perform calculations involving the idea of limiting reactants*

Let's consider making a chicken sandwich:

2  + 1  → 1 

But 1000 meat patties and 2500 slices of bread was given to the production line.

Consider:

Before (B)	2500	1000	0
Change (C)	-2000	-1000	+1000
After (A)	500	0	1000

When would the production stop? When the meat patties are used up.

Which ingredient determine the amount of sandwiches formed? Meat patties

Similarly, in a chemical reaction, the reactant that is used up **first** will determine the amount of product formed. And the reactant that is used up first is known as the **limiting reagent**.

Example 1:

If 8 g of copper was mixed with 4 g of sulfur, find the mass of copper(I) sulfide formed.

$$2\ Cu\ (s) \quad + \quad S\ (s) \rightarrow \quad Cu_2S\ (s)$$

amt of Cu = 8/64 = 0.125 mol

amt of S = 4/32 = 0.125 mol

Assume Cu is the limiting reagent, 0.125 mol of Cu will be used up and 0.0625 mol of S will be required for reaction.

Therefore S is in excess (since there is 0.125 mol initially) and Cu is the limiting reagent.

$2Cu \equiv Cu_2S$

amt of Cu_2S = 0.125 / 2 = 0.0625 mol

mass of Cu_2S = 0.0625 x 160 = 10.0g

Example 2:

A sample of 50 cm^3 of propane was burnt in 50 cm^3 of oxygen gas.

$$C_3H_8\ (g) + 5\ O_2\ (g) \rightarrow 3\ CO_2\ (g) + 4\ H_2O\ (g)$$

(a) Find the volume of carbon dioxide gas produced.

(b) Find the volume of excess reactant left.

Assume that propane is the limiting reagent, 50 cm^3 of propane will be used up and 250 cm^3 of oxygen will be required.

Therefore oxygen is the limiting reagent instead (as there is only 50cm^3 initially) and propane is in excess.

(a) $3\ CO_2 \equiv 5\ O_2$

Volume of carbon dioxide = $\frac{3}{5}x50 = 30cm^3$

(b) 50 cm^3 of oxygen will used up 10 cm^3 of propane.

Volume of propane left = 50 – 10 = 40 cm^3.

Check Point 8

1 If 7.2 g of iron(II) oxide was mixed with 0.4 dm^3 of oxygen gas, according to this equation:

$$4\ FeO + O_2 \rightarrow 2\ Fe_2O_3$$

(a) find the mass of iron(III) oxide formed.

(b) find the mass of excess reactant left.

Amount of FeO = 7.2/(56+16) = 0.1 mol

Amount of oxygen = 0.4/24 = 0.01667 mol

Assume that oxygen is the limiting reagent, 0.01667 mol of oxygen will be used up and 0.006667 mol of FeO will be required. FeO will be in excess (since there is 0.1 mol initially) and oxygen is the limiting reagent.

$2\ Fe_2O_3 \equiv O_2$

Amt of Fe_2O_3 = 0.01667 x 2 = 0.03334 mol

(a) mass of iron(III) oxide formed = 0.03334 x (56+56+16+16+16) = 5.33g

(b) amt of excess FeO = 0.1 – 0.06667 = 0.03333 mol

mass of excess FeO = 0.03333 x 72 = 2.40 g

2 A mixture of 30 cm^3 of hydrogen and 20 cm^3 of oxygen is exploded, find:

(a) the mass of steam formed.

(b) the volume of excess reactant gas left.

$$2H_2 + O_2 \rightarrow 2H_2O$$

Assume that hydrogen is the limiting reagent, 30 cm^3 of hydrogen will be used up and 15 cm^3 of oxygen will be required. Oxygen is in excess (as there is 20 cm^3 initially) and hydrogen is the limiting reagent.

$2H_2 \equiv 2H_2O$

(a) mass of stream = 30 cm^3

(b) vol of excess reactant = 20-15 = 5 cm^3

3 Ammonia burns in oxygen according to the following equation:

$$4\ NH_3\ (g) + 3O_2\ (g) \rightarrow 2\ N_2\ (g) + 6\ H_2O\ (l)$$

If 40.0 cm^3 of ammonia is burnt in 50.0 cm^3 of oxygen, what will be the volume of the gases after the combustion, assuming all the volumes are measured at r.t.p.?

Assume that ammonia is the limiting reagent, 40 cm^3 of ammonia will be used up and 30 cm^3 of oxygen will be required. Oxygen is in excess (as there is 50 cm^3 initially) and ammonia is the limiting reagent.

40cm^3 of ammonia will react with <u>30cm^3 of oxygen and produced 20cm^3 of nitrogen gas</u>.

N Percentage Yield and Purity

At the end of this section, you should be able to:-

- *Understand that percentage yield in an experiment cannot be 100%*
- *Calculate percentage purity based on theory and experimental results*

Percentage Yield

- The amount of product formed in a reaction is known as the **yield**.
- The **theoretical yield** of a reaction is the calculated amount of products that would be obtained if the reaction is completed.
- The **actual yield** is the amount of pure products that is actually produced in the experiment.
- The actual yield is usually **lesser** the theoretical yield.
- Formula:

$$\text{Percentage yield} = \frac{actual\ yield}{theoreteical\ yield} x100\%$$

Example 1:

When 1.92 g of magnesium was heated in excess oxygen, 3.0 g of magnesium oxide was obtained. Calculate the percentage yield of magnesium oxide.

$2\ Mg + O_2 \rightarrow 2MgO$

Amount of Mg = 1.92 / 24 = 0.08 mol

Mg≡MgO

Theoretical amount of MgO formed = 0.08mol

Theoretical mass of MgO formed = 0.08 x 40 = 3.2g (theoretical yield)

% yield = $\frac{3}{3.2} x100\% = 93.8\%$

Percentage Purity

- gives the indication of the amount of substance present in a mixture (impure substance)

$$\text{Percentage purity} = \frac{mass\ of\ pure\ substance\ in\ sample}{mass\ of\ sample} x100\%$$

<u>Example 2:</u>

3.2 g of copper was heated in air. 3.8 g of copper(II) oxide was formed. What is the percentage purity of copper?

$2Cu + O_2 \rightarrow 2CuO$

Amount of CuO = $\frac{3.8}{80} = 0.0475\ mol$

Cu ≡ CuO

Amount of pure Cu = $0.0475\ mol$

Mass of pure Cu = 0.0475 x 56 = 3.04g

% purity = $\frac{3.04}{3.2} x100\% = 95.0\%$

Check Point 9

1 In an experiment, 30 g of ethanoic acid (CH_3COOH) were obtained from the oxidation of 69 g of ethanol (C_2H_5OH). Calculate the percentage yield in the reaction.

$$C_2H_5OH + 2\,[O] \rightarrow CH_3COOH + H_2O$$

Amount of ethanol = 69/46 = 1.5 mol

$C_2H_5OH \equiv CH_3COOH$

Theoretical amount of CH_3COOH = 1.5 mol

Theoretical mass of CH_3COOH = 1.5 x 60 = 90 g (theoretical yield)

% yield = 30/90 x 100% = 33.3%

2 3.21 g sample of impure copper(II) carbonate was reacted with excess hydrochloric acid. It was found that 480 cm^3 of carbon dioxide gas measured at r.t.p. was given off. What is the percentage purity of the copper(II) carbonate in the given sample?

$CuCO_3 + 2HCl \rightarrow CuCl_2 + CO_2 + H_2O$

Amount of CO_2 = 0.48/24 = 0.02 mol

$CuCO_3 \equiv CO_2$

Amount of pure $CuCO_3$ = 0.02 mol

Mass of pure $CuCO_3$ = 0.02 x 124 = 2.48 g

% purity = $\frac{2.48}{3.21}x100\%$ = 77.3%

O Titration Calculations

At the end of this section, you should be able to:-

- *Perform calculation based on data obtained from titration results.*

Example 1:

Aqueous sodium sulfate can be prepared by titrating dilute sulfuric acid with aqueous sodium hydroxide. The equation for the reaction is:

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

In the titration, 25.0 cm^3 of 2.0 mol/dm^3 sulfuric acid was used.

(a) Calculate the volume of 0.5 mol/dm^3 sodium hydroxide used in the reaction.

(b) Calculate the mass in grams, of sodium sulfate produced in the reaction.

Amount of sulfuric acid = $\frac{25}{1000}x2 = 0.05\ mol$

H_2SO_4 (aq) ≡ 2NaOH

Amount of NaOH = 0.05 x 2 = 0.1 mol

(a) Volume of NaOH = $\frac{0.1}{0.5} = 0.200\ dm^3$

(b) $H_2SO_4 \equiv Na_2SO_4$

Amt of Na_2SO_4 = 0.05 mol

Mass of Na_2SO_4 = 0.05 x 142 = 7.10g

Example 2:

A concentrated acid has a concentration of 12 mol/ dm^3. What volume of concentrated acid would be needed to make 1 dm^3 of dilute acid with a concentration of 1 mol/ dm^3.

T to introduce $\boxed{C_1V_1 = C_2V_2}$

$12 \times V_? = 1 \times 1$
$V_? = 1/12 = 0.0833 dm^3$

Formulae, Stoichiometry, Mole Concept
Practice Worksheet 9D

Answer the following questions on writing paper.

1 28.0 g of nitrogen reacted with 8.0 g of hydrogen to form 5.1 g of ammonia.

(a) Which reactant is in excess?

(b) What is the percentage yield of ammonia?

Amount of nitrogen = 28/28 = 1 mol

Amount of hydrogen = 8/2 = 4 mol

$N_2 + 3H_2 \rightarrow 2NH_3$

Assume nitrogen is the limiting reagent, 1 mole of nitrogen will be used up and 3 moles of hydrogen will be required. Hydrogen will be in excess (as there is 4 moles initially). Therefore nitrogen is the limiting reagent.

(a) hydrogen is in excess.

(b) $N_2 \equiv 2NH_3$

Theoretical amt of ammonia produced = 2 mol

Theoretical mass of ammonia produced = 2 x 17 = 34g (theoretical yield)

%yield = $\frac{5.1}{34} x\ 100\% = 15.0\%$

2 In the fermentation process, glucose ($C_6H_{12}O_6$) is converted into ethanol (C_2H_5OH) by yeast in the following reaction:

$$C_6H_{12}O_6 \rightarrow 2\ C_2H_5OH + 2\ CO_2$$

Find the mass of ethanol that can be obtained from 45 g of glucose assuming <u>60% conversion of glucose</u> into ethanol.

Means 60% of 45g is used for the conversion

Mass of glucose used = 60/100 x 45 = 27 g

Amount of glucose used = $\frac{27}{180} = 0.15\ mol$

$C_6H_{12}O_6 \equiv 2\ C_2H_5OH$

Amount of ethanol formed = 0.15 x 2 = 0.3 mol

Mass of ethanol formed = 0.3 x 46 = 13.8 g

3 From 200 tonnes of propene (C_3H_6), the manufacturer obtained 150 tonnes of propan-2-ol (C_3H_8O). Calculate the percentage yield.

$$C_3H_6 + H_2O \rightarrow C_3H_8O$$

1 tonne = 1000kg = $1\text{x}10^6$ g

Amt of propene = 200000 / 42 = 4761.9 mol

$$C_3H_6 \equiv C_3H_8O$$

Theoretical Amt of propanol = 4761.9 mol

Theoretical mass of propanol = 4761.9 x 60 = 285714 g

% yield = $\frac{150000}{285714} x\ 100\% = 35.0\%$

4 A 10 g sample of impure ammonium sulfate was heated with excess sodium hydroxide. Ammonia gas and water were produced. The ammonia gas liberated was measured to be 2.4 dm^3. Calculate the percentage purity of ammonium sulfate in the sample.

$(NH_4)_2SO_4 + 2NaOH \rightarrow 2NH_3 + 2H_2O + Na_2SO_4$

Amt of ammonia = 2.4 / 24 = 0.1 mol

$(NH_4)_2SO_4 \equiv 2NH_3$

Amt of pure $(NH_4)_2SO_4$ = 0.1 / 2 = 0.05 mol

Mass of pure $(NH_4)_2SO_4$ = 0.05 x 132 = 6.6g

% purity = 6.6 / 10 x 100% = 66.0%

5 What volume (in cm^3) of 0.2 mol/dm^3 hydrochloric acid is required to react completely with 25cm^3 of 0.2 mol/dm^3 aqueous sodium carbonate?

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

Amount of Na_2CO_3 = $\frac{25}{1000} x 0.2 = 0.005\ mol$

$2HCl \equiv Na_2CO_3$

Amt of HCl = 0.005 x 2 = 0.01 mol

Volume of HCl = 0.01 / 0.2 = 0.05 dm^3 = 50.0 cm^3

6 A 300mg sample of acetylsalicylic acid (Mr = 180) is dissolved in some water. This acetylsalicylic acid solution requires 16.70 cm^3 of 0.100 mol/dm^3 aqueous sodium hydroxide for complete neutralization.

Calculate the number of moles of acetylsalicylic acid required to react with one mole of aqueous sodium hydroxide.

Amount of NaOH = 16.7/1000 x 0.1 = 0.00167 mol

Amount of acid = 300 x 10^{-3} / 180 = 0.001666 mol

0.001666 mol of acid reacts with 0.00167 mol of NaOH.

Therefore 1 mol of acid reacts with 1 mol of NaOH.

7. 18.0 cm^3 of 0.2 mol/dm^3 H_2SO_4 reacts with 24.0 cm^3 of KOH solution. Calculate the conc. of KOH in
   (a) mol/dm^3
   (b) g/dm^3

   (a) $H_2SO_4 + 2KOH \rightarrow K_2SO_4 + 2H_2O$
   Amount of H_2SO_4 = 18/1000 x 0.2 = 0.0036 mol
   $H_2SO_4 \equiv 2KOH$
   Amount of KOH =0.0036 x 2 = 0.00720 mol
   Concentration of KOH = 0.0072 x $\frac{1000}{24} = 0.300\ mol/dm^3$
   (b)
   Concentration of KOH in g/dm^3 = 0.300 x $(39 + 16 + 1) = 16.8\ mol/dm^3$