Stoichiometry and Mole Concept (Part 1)

Learning Objectives:

- define relative atomic mass, A_r
- define relative molecular mass M_r, and calculate relative molecular mass (and relative formula mass) as the sum of relative atomic masses
- calculate the percentage mass of an element in a compound when given appropriate information
- calculate empirical and molecular formulae from relevant data
- calculate stoichiometric reacting masses and volumes of gases (one mole of gas occupies 24.0 dm^3 at room temperature and pressure); calculations involving the idea of limiting reactants may be set
- apply the concept of solution concentration (in mol/dm^3 or g/dm^3) to process the results of volumetric experiments and to solve simple problems
- calculate percentage yield and percentage purity

1.1 Relative Atomic Mass, A_r

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

$$\textbf{Relative atomic mass, } A_r = \frac{\textbf{average mass of one atom of the element}}{\frac{1}{12}\textbf{ mass of an atom of carbon-12}}$$

- Relative atomic mass is a ratio and has no unit.
- The relative atomic masses of some elements are given below:

Element	Relative atomic mass (A_r)
hydrogen	1.0
carbon	12.0
oxygen	16.0
chlorine	35.5
magnesium	24.3

Some A_r values are not whole numbers

- Why are some A_r values not whole numbers?

Example: Why is A_r of chlorine 35.5?

The A_r of an element has to take into account the masses of all the isotopes of the element and their relative abundances.

Chlorine exists as a mixture of two isotopes, 75.0% of chlorine-35 and 25.0% of chlorine-37.

Hence, the relative atomic mass of chlorine $= (0.75 \times 35) + (0.25 \times 37)$
$= 26.25 + 9.25$

$= 35.5$

Sum of the A_r of all the atoms in a molecule

1.2 Relative Molecular Mass and Relative Formula Mass, M_r

- Definition: The ***relative molecular mass*** **(M_r) *of a molecule is the*** average **mass of one *molecule of a substance when compared with*** $\frac{1}{12}$ ***the mass of one atom of carbon-12.***

Relative molecular mass, M_r = $\dfrac{\text{average mass of one molecule of an element or a compound}}{\frac{1}{12}\text{ mass of an atom of carbon-12}}$

1 d.p
- Relative molecular mass is a ratio and has no unit.
- To find the M_r of a molecule, we just ***add up the relative atomic masses of all atoms*** in the molecule.
- **Ionic** compounds such as sodium chloride and magnesium nitrate **consist of ions** and **not molecules**. Hence, for ionic compounds, the term ***relative*** formula ***mass*** (M_r) is used instead of relative molecular mass.

1) Statement
2) Working
3) Answer to the correct precision
1 d.p

Example:
Find the relative molecular mass/relative formula mass of
(a) carbon dioxide (b) ammonia gas (c) sucrose, $C_{12}H_{22}O_{11}$ (d) $CuSO_4.5H_2O$

plus
water of crystallisation

(a) M_r of carbon dioxide = $(1 \times 12.0) + (2 \times 16.0)$
$= 44.0$

(b) M_r of ammonia gas = $(1 \times 14.0) + (3 \times 1.0)$
$= 17.0$

(c) M_r of $C_{12}H_{22}O_{11}$ = $(12 \times 12.0) + (22 \times 1.0) + (11 \times 16.0)$
$= 342.0$

(d) M_r of $CuSO_4.5H_2O$ = $(1 \times 63.5) + (1 \times 32.1) + (4 \times 16.0) + (5 \times 18.0)$

$= 249.6$

Quick Check 1

Homework

Calculate the relative molecular mass/relative formula mass of the following substances:

(a) H_2SO_4 $= (2 \times 1.0) + (1 \times 32.1) + (4 \times 16.0)$ $= 98.1$	(b) $Ca(NO_3)_2$ $= (1 \times 40.1) + 2[(1 \times 14.0) + (3 \times 16.0)]$ $= 164.1$
(c) $(NH_4)_2Cr_2O_7$ $= 2[(1 \times 14.0) + (4 \times 1.0)] + (2 \times 52.0) + (7 \times 16.0)$ $= 252$	(d) $Cu(NH_3)_4SO_4$ $= (1 \times 63.5) + 4[(1 \times 14.0) + (3 \times 1.0)] + (1 \times 32.1) + (4 \times 16.0)$ $= 227.60$
(e) CH_3COOH $= (1 \times 12.0) + (3 \times 1.0) + (1 \times 12.0) + (2 \times 16.0) + (1 \times 12.0)$ $= 71$	(f) $Na_2CO_3.10H_2O$ $= (2 \times 23) + (1 \times 12.0) + (3 \times 16.0) + 10[(2 \times 1.0) + (1 \times 16.0)]$ $= 286$
(g) $K_2SO_4.Al_2(SO_4)_3.24H_2O$ $= (2 \times 39.1) + (1 \times 32.1) + (4 \times 16.0) + (2 \times \quad) + 3[(1 \times 32.1) + (4 \times 16.0)] + 24[(2 \times 1.0) + (1 \times 6.0)]$ $= 2666.9$	

M_r of water = 18.0

2.1 Composition of Compounds by Mass

The percentage of an element in a compound by mass

$$= \frac{\text{No. of atoms of the element in the formula} \times A_r \text{ of the element}}{M_r \text{ of the compound}} \times 100\%$$

Example 1

What is the percentage by mass of hydrogen present in hydrogen peroxide, H_2O_2?

M_r of $H_2O_2 = (2 \times 1.0) + (2 \times 16.0) = 34.0$

Percentage by mass of hydrogen in H_2O_2

$= \frac{1.0 \times 2}{34.0} \times 100\%$

$= 5.88\%$ (3 s.f)

Example 2

What is the percentage of water in copper(II) sulfate crystals, $CuSO_4.5H_2O$?

Percentage by mass of water in $CuSO_4 \cdot 5H_2O$

$= \frac{5 \times 18.0}{63.5 + 32.1 + (4 \times 16.0) + (5 \times 18.0)} \times 100\%$ → M_r of $CuSO_4 \cdot 5H_2O$

$= \frac{5 \times 18.0}{249.6} \times 100\%$

$= 36.1\%$ (3 s.f)

Quick Check 2

1 Calculate the following:
(a) percentage of carbon in ethanoic acid, CH_3COOH → 180
(b) percentage of water in $Na_2CO_3.10H_2O$

2(a) Percentage of carbon in ethanoic acid

$= \frac{12.0 \times 2}{12.0 + (3 \times 1.0) + 12.0 + 16.0 + 16.0 + 1.0}$

$= \frac{24.0}{60.0} \times 100\%$

$= 40.0\%$ (3 s.f)

b) Percentage of water in $Na_2CO_3 \cdot 10H_2O$

$= \frac{18.0 \times 10}{(2 \times 23.0) + (12.0) + (3 \times 16.0) + (10 \times 18.0)} \times 100\%$

$= 62.9\%$ (3 s.f)

↳ penalised if not 3 s.f

2 An iron ore extraction involves the extraction of haematite, which contains Fe_2O_3, and magnetite, which contains Fe_3O_4. Which has greater percentage of iron by mass, Fe_2O_3 or Fe_3O_4?

% of Fe in Fe_2O_3 = $\frac{(2 \times 55.8)}{(2 \times 55.8) + (3 \times 16.0)} \times 100\%$

$= 69.9\%$ (3 s.f)

% of Fe in Fe_3O_4 = $\cancel{\frac{(3 \times 55.8)}{(3 \times 55.8) + 4 + (8)}}$ $\frac{(3 \times 55.8)}{(3 \times 55.8) + (4 \times 16.0)} \times 100\%$

$= 72.3\%$ (3 s.f)

Therefore, Fe_3O_4 contains greater % of Fe.

Note: What is the percentage of oxygen in Fe_3O_4? *Answer:* ~~29.~~ 27.7

Sum of percentages of individual elements in compound = 100 %

2.2 Calculating the Mass of an Element/Water of Crystallisation in a Compound

The mass of an element in a sample of a compound

$$= \frac{\text{No. of atoms of the element in the formula} \times A_r \text{ of the element}}{M_r \text{ of the compound}} \times \text{Mass of sample}$$

Example 1

Calculate the mass of oxygen in 32.0 g of copper(II) sulfate, $CuSO_4$.

Mass of oxygen(O) in $CuSO_4$ = $\frac{4 \times \text{Relative atomic mass of O}}{\text{Relative formula mass of } CuSO_4} \times$ mass of sample

$= \frac{4 \times 12.0}{63.5 + 32.1 + 4(16.0)} \times 32.0$

$= 12.8 \text{ g}$ (3s.f)

Example 2

Calculate the mass of water in 12.5 g of hydrated copper(II) sulfate, $CuSO_4.5H_2O$.

Mass of water(H_2O) in $CuSO_4.5H_2O$ = $\frac{5 \times \text{Relative molecular mass of } H_2O}{\text{Relative formula mass of } CuSO_4.5H_2O} \times$ mass of sample

$= \frac{5 \times 18.0}{249.6} \times 12.5\text{g}$

$= 4.51\text{g}$ (3s.f)

Quick Check 3

1 Calculate the mass of
(a) calcium in 25.0 g of calcium carbonate;
(b) chlorine in 27.0 g of copper(II) chloride;
(c) water in 54.1 g of $FeCl_3.6H_2O$.

a) Mass of calcium in $CaCO_3 = \frac{40.1}{40.1 + (16.0 \times 3) + 12.0} \times 25.0$

$= 10.0\text{g}$ (3s.f)

b) Mass of chlorine in $CuCl_2 = \frac{35.5 \times 2}{63.5 + (35.5 \times 2)} \times 27.0$

$= 14.3\text{ g}$ (3s.f)

c) Mass of water in $FeCl_3.6H_2O = \frac{6(18.0)}{55.8 + (35.5 \times 3) + 6(18.0)} \times 54.1$

$= 21.6\text{g}$ (3s.f)

2 A sample of iron chloride has the formula $FeCl_n$ where n is a whole number. If the percentage of chlorine in the compound is 65.5%, what is the value of n? $n = 3$

Percentage ~~Mass~~ of chlorine in $FeCl_n$ = 65.5%

Percentage of iron in $FeCl_n$ = 100% − 65.5%

= 34.5%

34.5% = 55.8

1% = 1.6174

100% = ~~1.61~~ 161.74

value of n = $\frac{161.74 - 55.8}{35.5}$

= 2.98 (2d.p)

3.1 The Mole and Avogadro's Number

- The mole is the counting unit used by chemists for very small particles such as atoms, molecules, ions and electrons.
- It is taken from the Latin word, "moles", meaning "heap" or "pile".
- **One mole of a substance contains the same number of particles as the number of atoms in 12 g of carbon-12.**
- The number of atoms in 12 g of the carbon-12 isotopes is known as Avogadro's constant/ Avogadro's number.

  Its value is 6.02×10^{23}.

- The symbol of mole is mol. The SI unit for the amount of substance is **mole**.
- One mole of any substance contains ***6.02 x 10^{23}* particles.**

Items	Unit used to count	Number of items in one unit
socks	pair	2
eggs	dozen	12
bottles	carton	24
papers	ream	500
atoms, molecules, ions, electrons	mole	6.02×10^{23}

Examples

(a) 1 mole of neon, Ne, contains 6.02×10^{23} neon atoms.

(b) 1 mole of hydrogen gas, H_2, contains
1 mole of hydrogen molecules or 6.02×10^{23} hydrogen molecules,
2 moles of hydrogen atoms or $2 \times (6.02 \times 10^{23})$ hydrogen atoms.

(c) 1 mole of methane gas, CH_4, contains
1 mole of methane molecules or 6.02×10^{23} methane molecules,
1 mole of carbon atoms or ~~XX~~ 6.02×10^{23} carbon atoms,
4 moles of hydrogen atoms or $4 \times 6.02 \times 10^{23}$ hydrogen atoms.

(d) 1 mole of sodium chloride, NaC*l*, contains
1 mole of formula units(NaC*l*) or 6.02×10^{23} formula units(NaC*l*),
1 mole of sodium ions or 6.02×10^{23} sodium ions,
1 mole of chloride ions or 6.02×10^{23} chloride ions.

(e) 1 mole of magnesium chloride, $MgCl_2$, contains
1 mole of formula units($MgCl_2$) or 6.02×10^{23} formula units($MgCl_2$),
1 mole of magnesium ions or 6.02×10^{23} magnesium ions,
2 moles of chloride ions or $2 \times 6.02 \times 10^{23}$ chloride ions.

- Since 1 mole of a substance contains 6.02 x 10^{23} particles, we can use a convenient formula to find the number of moles of particles in a substance.

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

Finalanswer: 3s.f

Example 1

How many moles of water are there in 3.00 x 10^{24} molecules of water?

Number of moles of water molecules

$$= \frac{\text{number of water molecules}}{6.02 \times 10^{23}}$$

$$= \frac{3.00 \times 10^{24}}{6.02 \times 10^{23}} = 4.98 \text{ mol}$$

Example 2

How many atoms are there in 0.300 mol of iron?

Number of iron atoms

= number of moles of iron atoms $\times 6.02 \times 10^{23}$

$= 0.300 \times 6.02 \times 10^{23}$

$= 1.81 \times 10^{23}$ //

Example 3

What is the number of (a) ammonia molecules; (b) nitrogen atoms; (c) hydrogen atoms in 0.250 moles of ammonia gas?

a) Number of ammonia molecules

$= 0.250 \times \cancel{4}\ 6.02 \times 10^{23}$

$= 1.51 \times 10^{23}$

b) Number of nitrogen atoms

$= 0.250 \times 6.02 \times 10^{23}$

$= 1.51 \times \cancel{2}\ 10^{23}$

c) Number of hydrogen atoms

$= 0.250 \times 5.02 \times 10^{23} \times 3$

$= 4.52 \times 10^{23}$ //

Quick Check 4

1 Calculate the number of moles of atoms in 2.40 x 10^{25} molecules of hydrogen gas, H_2.

Number of moles of atoms in 2.40×10^{25} molecules of H_2

$$= \frac{2 \times 2.40 \times 10^{25}}{6.02 \times 10^{23}}$$

$= 79.7$ mol

$Zn(NO_3)_2$

2 Calculate the number of (a) zinc ions and (b) nitrate ions in 0.500 moles of zinc nitrate.

a) Number of zinc ions: $0.500 \times 6.02 \times 10^{23}$

$= 3.01 \times 10^{23}$ //

b) Number of nitrate ions: $0.500 \times 6.02 \times 10^{23} \times 2$

$= 6.02 \times 10^{23}$ //

3 Given 0.250 moles of carbon dioxide, calculate

(a) the number of carbon dioxide molecules,

(b) the number of atoms that are present.

a) Number of CO_2 molecules

$= 0.250 \times 6.02 \times 10^{23}$

$= 1.505 \times 10^{23}$

$= 1.51 \times 10^{23}$ // (3 s.f)

b) Number of atoms that are present

$= 3 \times 1.505 \times 10^{23}$

$= 4.515 \times 10^{23}$

$= 4.52 \times 10^{23}$ //

3.2 Molar Mass

- ***Molar mass is the mass of*** one ***mole of any substance.***

(i) For substances consisting of atoms

The molar mass is the **relative atomic mass** of the element in grams per mole (g/mol).

E.g. Molar mass of carbon = 12.0 g/mol

Molar mass of sodium = 23.0g/mol

(ii) For substances consisting of molecules

The molar mass is the **relative molecular mass** of the substance in grams per mole.

E.g. Molar mass of oxygen gas = 16 x 2 = 32.0 g/mol

Molar mass of water = 18.0g/mol

(iii) For substances consisting of ions

The molar mass is the **relative formula mass** of the substance in grams per mole.

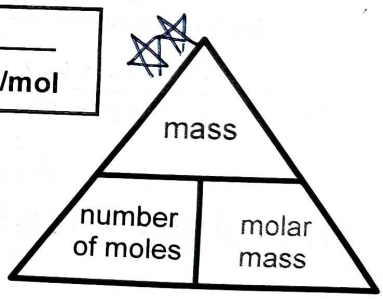
E.g. Molar mass of sodium chloride = 23.0 + 35.5 = 58.5 g/mol

(132.0g/mol (3s.f))

Molar mass of ammonium sulfate = 2[14.0+(4x1.0)]+32.1+(4x16.0) = 132.1g/mol

- If we know the mass of a substance, we can calculate the number of moles using the formula:

$$\textbf{Number of moles of substance} = \frac{\textbf{Mass of substance in g}}{\textbf{Molar mass of substance in g/mol}}$$



Example 1

Calculate the number of moles of atoms in 8.00 g of helium. → monatomic (Noble gas)

Number of moles of atoms in 8.00g of helium

$= \frac{8.00}{4.0 g/mol}$

$= 2.00 mol$

Example 2

Calculate the number of moles of molecules in 4.50 g of water. H_2O

Number of moles of ~~ato~~ molecules in 4.50g of water

$= \frac{4.50}{2(1.0)+16.0}$

$= 0.250 mol$

Example 3

Calculate the mass of 0.400 moles of iron atoms.

~~Name~~ Mass of 0.400 moles of iron atoms

$= 0.400 \times 55.8$

$= 22.32$

$= 22.3 g/mol$

Example 4 $\nearrow CaCl_2$

Calculate the number of chloride ions in 6.66 g of calcium chloride.

Number of moles of Cl^- in 6.66g of calcium chloride

$= \frac{6.66}{40.1+2(35.5)} \times 2$

$= 0.11989$ mol // (5s.f)

Number of Cl^- in 6.66g of $CaCl_2$

$= 0.11989 \times 6.02 \times 10^{23}$

Example 5 $= 7.22 \times 10^{22}$ // (3s.f)

0.250 mol of fructose has a mass of 45.0 g. What is the relative molecular mass of fructose?

Relative molecular mass of fructose

Molar mass of fructose $= \frac{45.0}{0.250}$

$= 180.0$ g/mol

Relative molecular mass of fructose $= 180.0$ ~~[illegible]~~ (No units)

Quick Check 5

1 Calculate the number of moles of atoms for each of the following:

(a) 32.0 g of oxygen (b) 8.00 g of calcium

(c) 3.00 x 10^{23} iodine atoms (d) 1.50 x 10^{24} carbon atoms

a) Number of moles of atoms for 32.0g of oxygen

$= \frac{32.0}{2 \times 16.0 \text{ g/mol}} \times 2$

$= 2.00$ mol // (3s.f)

b) Number of moles of atoms for 8.00g of calcium

$= \frac{8.00}{40.1 \text{ g/mol}}$

$= 0.200$ mol (3s.f)

c) Number of moles of atoms for 3.00×10^{23} iodine atoms

$= \frac{3.00 \times 10^{23}}{6.02 \times 10^{23}}$

$= 0.498$ mol //

d) Number of moles of atoms for 1.50×10^{24} carbon atoms

$= \frac{1.50 \times 10^{24}}{6.02 \times 10^{23}}$

$= 2.49$ mol (3s.f)

2 Calculate the number of moles of molecules for each of the following:

(a) 32.0 g of sulfur dioxide (b) 20.0 g of nitrogen gas

(c) 2.40 x 10^{23} molecules of hydrogen (d) 12.0 x 10^{24} molecules of ammonia

a) Number of moles of molecules for 32.0g of SO_2

$= \frac{32.0}{(32.1)+(32.0) \text{ g/mol}}$

$= 0.499$ mol // (3s.f)

b) Number of moles of molecules for 20.0g of N_2

$= \frac{20.0}{2(14.0) \text{ g/mol}}$

$= 0.714$ mol // (3s.f)

c) Number of moles of molecules for 2.40×10^{23} molecules of hydrogen

$= \frac{2.40 \times 10^{23}}{6.02 \times 10^{23}}$

$= 0.399$ mol // (3s.f)

d) Number of moles of molecules for ~~to~~ 12.0×10^{24} molecules of ammonia

$= \frac{12.0 \times 10^{24}}{6.02 \times 10^{23}}$

$= 19.9$ mol // (3s.f)

3 Calculate the mass of each of the following:
(a) 0.600 mole of magnesium chloride
(b) 2.00 moles of oxygen gas
(c) 6.00 x10^{23} atoms of copper
(d) 1.20 x 10^{24} molecules of chlorine gas

a) Mass of 0.600 mole of magnesium chloride
= (0.600)(24.3 + 35.5)
= 35.88
= 35.9g (3s.f)

b) Mass of 2.00 moles of oxygen gas
= (2.00)(16.0 x 2)
= 64g
= 64.0g (3s.f)

c) Mass of 6.00 x 10^{23} atoms of copper
= 6.00 x 10^{23} x (63.5)
= 3.81 x 10^{25}

d) Mass of ~~10~~ 1.20 x 10^{24} molecules of chlorine gas
= 1.20 x 10^{24} x (35.5x2)
= 85.2 x 10^{24} g

4 How many hydrogen chloride, HC*l* molecules would have a mass of 292 g?

Number of moles of molecules of HCl = $\frac{292}{1.0+35.5}$
= 8.00 mol (3s.f)

Number of molecules of HCl =

5 What is the number of (a) molecules; (b) atoms in 4.50 g of water?

a) Number of moles of molecules in 4.50g of water
= $\frac{4.50}{18.0}$
= 0.250 mol (3s.f)

Number of molecules in 4.50g of water
= ~~0.5~~ 0.250 x 6.02 x 10^{23}
= 1.505 x 10^{23} x 3
= 1.31 x 10^{23}

b) Number of moles of atoms in 4.50g of water
= ~~0.25 x 6.02 x 10^{23}~~
~~= 1.~~
= 1.505 x 10^{23} x 3
= 4.52 x 10^{23} (3s.f)

6 How many ions are there in 60.0 g of magnesium chloride?

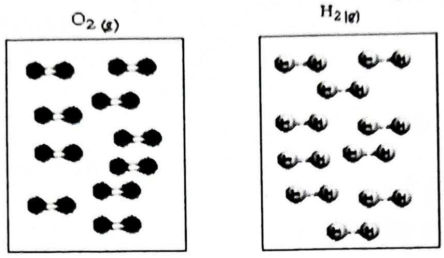
Number of moles of ions in 60.0g of MgCl
= $\frac{60.0}{24.3+(35.5\times2)}$
= 0.62959 mol

Number of ions = 0.62959 mol

Please do not leave ans in fractions

3.3 Molar Volume of Gases

- ***Avogadro's Law*** states that equal **volumes** of all gases, under the same **temperature** and **pressure** contain the **same** number of particles.
  → will not give same mass
  → Different molar mass



Equal Volumes and the Same Temperature and Pressure

- ***Molar volume*** of a gas is the volume occupied by one mole of any gas.
- At **room conditions** (i.e. **20°C** and **1 atmospheric pressure**), one mole of any gas at r.t.p. has a volume of **24.0** dm^3 or **24000** cm^3. This volume is called **molar volume** of a gas at r.t.p.
- Examples: Molar volume of oxygen gas at r.t.p. = **24.0 dm^3**
  Molar volume of helium gas at r.t.p. = **24.0 dm^3**
  Molar volume of carbon dioxide at r.t.p. = **24.0 dm^3**
- The conditions of temperature and pressure must be specified.
  - If temperature decreases, what happens to the molar volume of gas?
    The molar volume decreases.
  - At standard temperature (0 °C) and pressure (1 bar) (s.t.p.), the molar volume of a gas is **22.7** dm^3.
- Since at r.t.p., one mole of a gas would occupy a volume of 24.0 dm^3, we can use a convenient formula to find the number of moles of a gas.

$$\textbf{Number of moles of a gas} = \frac{\textbf{Volume of gas}}{\textbf{Molar volume of gas}}$$

← 24.0 dm^3 at r.t.p.

Example 1
Calculate the number of moles of carbon dioxide in 3.00 dm^3 of the gas at r.t.p.

Number of moles in carbon dioxide in 3.00 dm^3 of CO_2 at r.t.p
$= \frac{3.00}{24.0}$
$= 0.125$ mol

Example 2
Calculate the volume of 0.550 moles of methane gas at r.t.p.

Volume of 0.550 moles of methane gas at r.t.p
$= 0.550 \times 24.0$
$= 13.2\ dm^3$ (3s.f)

Example 3
Calculate the volume of 4.50 g of oxygen gas at r.t.p.

Volume of 4.50g of oxygen gas at r.t.p
$= \frac{4.50}{2(16.0)} \times 24.0\ dm^3$
$= 3.38\ dm^3$ (3s.f)

No. of moles of $O_2 = \frac{4.50}{2(16.0)}$
$= 0.14063$ mol
Volume of $O_2 = 0.14063 \times 24.0$
$= 3.38\ dm^3$

Example 4

What is the mass of 12.0 dm^3 of carbon dioxide at r.t.p.?

Number of moles of CO_2 $= \frac{12.0}{24.0}$

$= 0.500 mol$

Mass of CO_2 = no. of moles × molar mass

$= 0.500 \times [12.0 + (2 \times 16.0)]$

$= 0.500 \times 44.0$

$= 22.0 g$

Quick Check 6

1 Calculate the volume of the following gases at r.t.p.

(a) 0.200 mol of oxygen,

(b) 3.00 mol of hydrogen

Q1(a) Volume of 0.200 mol of oxygen at r.t.p

$= 0.200 \times 24.0$

$= 4.80 dm^3$

b) Volume of 3.00 mol of hydrogen at r.t.p

$= 3.00 \times 24.0$

$= 72.0 dm^3$

2 Calculate the number of moles of molecules of the following gases:

(a) 12.0 dm^3 of nitrogen at r.t.p.,

(b) 150 cm^3 of carbon monoxide at r.t.p.

Q2

a) Number of moles of molecules of 12.0 dm^3 of nitrogen at r.t.p

$= \frac{12.0}{24.0}$

$= 0.5 mol$

b) $150 cm^3 = 0.15 dm^3$

Number of moles of molecules of 150 cm^3 of carbon monoxide at r.t.p

$= \frac{0.15}{24.0}$

$= 0.00625 mol$

3 Calculate the mass of the following gases at r.t.p.

(a) 12.0 dm^3 of hydrogen,

(b) 480 cm^3 of carbon dioxide

Q3a) Number of moles of hydrogen

$= \frac{12.0}{24.0}$

$= 0.500 mol$

Mass of hydrogen $= 0.500 \times [(1.0) \times 2]$

$= 1.00 g$ (3s.f)

b) Number of moles of CO_2

$= \frac{480}{24.0}$

$= 20.0 mol$

Mass of carbon dioxide $= 20.0 \times [12.0 + 2(16.0)]$

$= 880 g$

$(NH_4)_2$ ...

4 Calculate the volume of the following gases at r.t.p.
(a) 14.0 g of nitrogen,
(b) 3.20 g of sulfur dioxide, SO_2
(c) 1.80×10^{23} molecules of chlorine gas

a) Number of moles of nitrogen

$= \frac{14.0}{2(14.0 g/mol)}$

$= 0.500$ mol (3s.f)

Volume of 14.0g of nitrogen at r.t.p

$= 0.500 \times 24.0$

$= 12\ dm^3$ //

b) Number of moles of sulfur dioxide

$= \frac{3.20}{32.1 + 2(16.0) g/mol}$

$= 0.049922$ (5s.f)

Volume of 3.20g of sulfur dioxide at r.t.p

$= 0.049922 \times 24$

$= 1.20\ dm^3$ (3s.f)

c) Number of mols of atoms for 1.80×10^{23} molecules of chlorine gas

$= \frac{1.80 \times 10^{23}}{6.02 \times 10^{23}}$

$= 0.29900$ mol (5s.f)

Volume of 1.80×10^{23} molecules of chlorine gas at r.t.p

$= 0.29900 \times 24.0$

$= 7.18\ dm^3$ //

5 A sample of carbon monoxide gas has a volume of 6.00 dm^3 at r.t.p.
Calculate
(a) the number of moles of gas,
(b) the number of molecules, → mol $\times (6.02 \times 10^{23})$
(c) the number of atoms,
(d) the mass of the carbon monoxide sample

a) Number of moles of gas in sample

$= \frac{6.00}{24.0}$

$= 0.25$ mol //

b) Number of molecules in sample

$= 0.25 \times 6.02 \times 10^{23}$

$= 1.505 \times 10^{23}$

$= 1.51 \times 10^{23}$ // (3s.f)

c) Number of atoms in sample

$= 1.505 \times 10^{23} \times 2$

$= 3.01 \times 10^{23}$ // (3s.f)

d) Mass of the carbon monoxide sample

$= 0.25 \times (12.0 + 16.0)$

$= 7.00$ g // (3s.f)

4. Empirical and Molecular Formulae

Chemical Formula

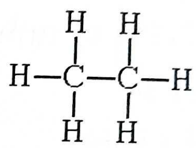
- **Molecular Formula** (shows the exact number of atoms of each element in a compound)
- **Empirical Formula** (shows the simplest ratio of the different types of atoms in a compound)
- **Structural Formula** (shows how the atoms are joined together in a compound)

Example : **Ethane**

Molecular formula: **C_2H_6**

Empirical formula: **CH_3**

Structural formula:



Compound	Molecular Formula	Empirical Formula
Water	H_2O	H_2O
Methane	CH_4	CH_4
Hydrogen Peroxide	H_2O_2	HO
Glucose	$C_6H_{12}O_6$	CH_2O
Ethene	C_2H_4	CH_2
Propene	C_3H_6	CH_2

Water, Methane: Molecular formula same as its empirical formula

Hydrogen Peroxide, Glucose: Molecular formula is a multiple of its empirical formula

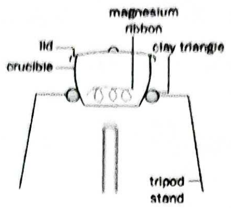
Ethene, Propene: Different compounds, same empirical formula

4. Calculating the Empirical Formula from Masses of Elements that are combined together

- The empirical formula of a compound can be found from the **masses** of elements that are combined together.
- The **number of moles** of atoms that are combined together are determined.
- The empirical formula is obtained using the simplest **mole ratio** of the atoms.

Exact : (3s.f)
Non-exact : (5s.f)

Example 1
To determine the empirical formula of magnesium oxide, magnesium ribbon is heated in a crucible as shown below.



The following results are obtained:

Mass of crucible + lid	= 26.52 g
Mass of crucible + lid + magnesium	= 27.72 g
Mass of crucible + lid + magnesium oxide	= 28.52 g

Mass of magnesium burnt = 27.72 - 26.52 = 1.20 g
Mass of oxygen that combined with magnesium = 28.52 - 27.72 = 0.80g

	Magnesium (Mg)	Oxygen (O)
Step 1 Write down the **mass** of each element (g)	1.20	0.80
Step 2 Write down the **molar mass** of each element (g/mol)	24.3	16.0
Step 3 Divide each mass by its molar mass to obtain the **number of moles of** each element (mol)	0.049383 (5s.f)	0.0500 (3s.f)
Step 4 Divide each no. of moles by the **smallest number**	$\frac{0.049383}{0.049383} = 1$	$\frac{0.0500}{0.049383} = 1$ (3s.f)

Empirical formula of magnesium oxide is MgO

Example 2
30.0 g of aluminium sulfide contains 19.2 g of sulfur. Find the empirical formula of the compound.

Element	Aluminium (Al)	Sulfide (S)
Mass (g)	30.0g - 19.2g = 10.8g	19.2
Molar mass (g/mol)	27.0	32.1
No. of moles (mol)	$\frac{10.8}{27.0} = 0.400$	$\frac{19.2}{32.1} = 0.59813$
Simplest ratio	$\frac{0.400}{0.400} = 1$ $1 \times 2 = 2$	$\frac{0.59813}{0.400} \approx 1.5$ $1.5 \times 2 = 3$

Empirical formula of the compound is Al_2S_3

4. Calculating the Molecular Formula from Empirical Formula

The molecular formula can be calculated from the **empirical formula** and **relative molecular mass**.

Example 1

Propene has the empirical formula CH_2. The relative molecular mass of propene is 42.0. Find the molecular formula of propene.

Let the molecular formula of propene be $(CH_2)_n$

M_r of $(CH_2)_n = 42.0$

$[12.0 + (1.0 \times 2)] \times n = 42.0$

$14.0n = 42.0$

$n = 42.0 \div 14.0$

$= 3$

Therefore, molecular formula of propene is C_3H_6

Example 2

Caffeine is a compound found in coffee and tea. The percentage composition of caffeine is 49.5% carbon, 5.10% hydrogen, 16.5% oxygen and 28.9% nitrogen.

The relative molecular mass of caffeine is 194. Determine the
(a) empirical formula
(b) molecular formula

- When percentage composition of the compound is given instead of masses of the elements, we assume 100 g of the compound is analysed.
- Each percentage then gives the mass of the element in grams in terms of 100 g of the compound

Let the mass of caffeine be 100g

Element	C	H	O	N
Mass in 100g (g)	49.5	5.10	16.5	28.9
Molar mass (g/mol)	12.0	1.0	16.0	14.0
No. of moles (mol)	$\frac{49.5}{12.0} = 4.125$	$\frac{5.10}{1.0} = 5.10$	$\frac{16.5}{16.0} = 1.0313$	$\frac{28.9}{14.0} = 2.0643$
Simplest ratio	$\frac{4.125}{1.0313} \approx 4$	$\frac{5.10}{1.0313} \approx 5$	$\frac{1.0313}{1.0313} = 1$	$\frac{2.0643}{1.0313} \approx 2$

(a) Empirical formula of caffeine is $C_4H_5ON_2$.

(b) Let the molecular formula of caffeine be $(C_4H_5ON_2)_n$

M_r of $(C_4H_5ON_2)_n = 194$

$[(12.0 \times 4) + (1.0 \times 5) + 16.0 + (14.0 \times 2)]_n = 194$

$97.0n = 194$

$n = 194 \div 97.0$

$= 2$ //

Therefore, molecular formula of caffeine is $C_8H_{10}O_2N_4$

Quick Check 7

1 Find the empirical formulae of the following:
(a) a compound consisting of 3.50 g of nitrogen combined with 8.00 g of oxygen;
(b) a compound consisting of 5.37 g of iron combined with 4.63 g sulfur.

a)

Element	Nitrogen	Oxygen
Mass (g)	3.50	8.00
Molar mass (g/mol)	14.0	16.0
No. of moles (mol)	0.25	0.5
Simplest Ratio	$\frac{0.25}{0.25} = 1$	$\frac{0.5}{0.25} = 2$

The empirical formula of the compound is NO_2 //

b)

Element	iron	sulfur
Mass (g)	5.37	4.63
Molar Mass (g/mol)	55.8	32.1
No. of Moles (mol)	0.096237	0.14424
Simplest Ratio	$\frac{0.096237}{0.096237} = 1$	$\frac{0.14424}{0.096237} = 1.5$

1 : 1.5 = 2 : 3

The empirical formula of compound is Fe_2S_3 //

2 Find the empirical formulae of the following:
(a) a compound with composition 46.7% silicon and 53.3% oxygen by mass;
(b) a compound with composition 10.0% carbon, 0.84% hydrogen and 89.1% chlorine by mass.

a)

element	silicon	oxygen
Percentage	46.7	53.3
Mass in 100g	~~49~~ 46.7	53.3
Molar mass (g/mol)	28.1	32.0
No. of Moles (mol)	$\frac{46.7}{28.1}$ = 1.6619 (5s.f)	$\frac{53.3}{32.0}$ = 1.6656 (5s.f)
Simplest Ratio	$\frac{1.6619}{1.6619} = 1$	$\frac{1.6656}{1.6619} = 1.0022$ (5s.f)

Ratio = 1 : 1
SiO_2

Empirical formulae = SiO_2 //

b)

Element	Carbon	Hydrogen	Chlorine
Mass in 100g	10.0	0.84	89.1
Molar mass (g/mol)	12.0	2.0	71
No. of moles (mol)	$\frac{10.0}{12.0}$ = 0.8333	0.0420	1.2549
Simplest Ratio	$\frac{0.8333}{0.0420}$ = 19.840 = 20	$\frac{0.0420}{0.0420} = 1$ = 1 × 2 = 2	$\frac{1.2549}{0.0420}$ = 29.879 = 30

	Carbon	Hydrogen	Chlorine
Mass in 100g	10.0	0.84	89.1
Molar Mass (g/mol)	12.0	1.0	71
No. of Moles	0.8333	0.84	1.2549
Simplest Ratio	$\frac{0.8333}{0.8333} = 1$	$\frac{0.84}{0.8333}$ ≈ 1	$\frac{1.2549}{0.8333} = 1.5$

1 : 1 : 1.5
= 2 : 2 : 3

∴ formula = $C_2H_2Cl_3$ //

3 A carbohydrate contains 54.5% carbon, 9.1% hydrogen, the balance being oxygen.

(a) Find the empirical formula of the carbohydrate.

(b) If the relative molecular mass of the carbohydrate is 88, what is its molecular formula?

Percentage of oxygen in the carbohydrate

$= 100 - 54.5 - 9.1$

$= 36.4\%$

$C_4H_8O \Rightarrow C_2H_4O$

let mass of carbohydrate be 100g

Element	carbon	hydrogen	oxygen
mass in 100g	54.5	9.1	36.4
molar mass (g/mol)	12.0	1.0	32.0
No. of moles (mol)	$\frac{54.5}{12.0} = 4.54167$	9.1	$\frac{36.4}{32.0} = 1.1375$
Simplest Ratio	$\frac{4.54167}{1.1375} = 3.9927$ $= 4$	$\frac{9.1}{1.1375} = 8$	$\frac{1.1375}{1.1375} = 1$

a) Empirical formula of carbohydrate: C_2H_4O //

b) let the molecular formula be $(C_2H_4O)_n$

M_r of $(C_2H_4O)n = 88$

$[2(12.0) + 4(1.0) + 32.0]n = 88$

$60n = 88$

$n = 1.5$ (nearest round)

$n = 2$

∴ Molecular formula $= C_4H_8O_2$ //

4 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. SiO_2

(b) If the mass of 1 mole of the oxide of silicon is 60.0 g, what is its molecular formula? SiO_2

a) Mass of silicon $= 15.48g - 15.20g$

$= 0.28g$

Mass of oxygen reacted with silicon $= 15.80g - 15.48g$

$= 0.32g$

Element	Si	O
mass	0.28	0.32
Molar mass	28.1	16.0
No. of moles	$\frac{0.28}{28.1} = 0.0099644$ $= 0.01$	$\frac{0.32}{16.0} = 0.02$
Simplest Ratio	$\frac{0.01}{0.01} = 1$	$\frac{0.02}{0.01} = 2$

ANSWERS TO QUICK CHECK

	Quick Check 1 (pg 2)		**Quick Check 5 (pg 8 - 9)**
(a)	98.1	3(a)	57.2 g
(b)	164.1	3(b)	64.0 g
(c)	252.0	3(c)	63.3 g
(d)	227.6	3(d)	142 g
(e)	60.0	4	4.82×10^{24}
(f)	286.0	5(a)	1.51×10^{23}
(g)	948.6	5(b)	4.52×10^{23}
		6	1.14×10^{24}
	Quick Check 2 (pg 3)		
1(a)	40.0 %		
1(b)	62.9 %		**Quick Check 6 (pg 11 - 12)**
2	69.9 % ; 72.3% ; Fe_3O_4 ; 27.7%; 100%	1(a)	4.80 dm^3
		1(b)	72.0 dm^3
	Quick Check 3 (pg 4)	2(a)	0.500 mol
1(a)	10.0 g	2(b)	0.00625 mol
1(b)	14.3 g	3(a)	1.00 g
1(c)	21.6 g	3(b)	0.880 g
2	3	4(a)	12.0 dm^3
		4(b)	1.20 dm^3
	Quick Check 4 (pg 6)	4(c)	7.18 dm^3
1	79.7 mol	5(a)	0.250 mol
2(a)	3.01×10^{23}	5(b)	1.51×10^{23}
2(b)	6.02×10^{23}	5(c)	3.01×10^{23}
3(a)	1.51×10^{23}	5(d)	7.00 g
3(b)	4.52×10^{23}		
	Quick Check 5 (pg 8 - 9)		**Quick Check 7 (pg 16 - 17)**
1(a)	2.00 mol	1(a)	NO_2
1(b)	0.200 mol	1(b)	Fe_2S_3
1(c)	0.498 mol	2(a)	SiO_2
1(d)	2.49 mol	2(b)	$CHCl_3$
2(a)	0.499 mol	3(a)	C_2H_4O
2(b)	0.714 mol	3(b)	$C_4H_8O_2$
2(c)	0.399 mol	4(a)	SiO_2
2(d)	19.9 mol	4(b)	SiO_2

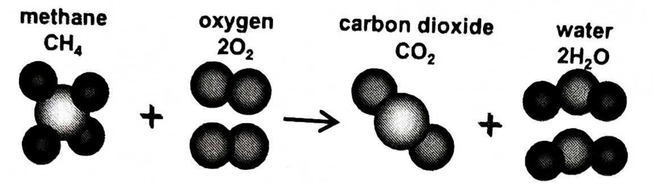
Stoichiometry and Mole Concept (Part 2)

5.1 Calculations from Equations

- A balanced chemical equation provides information on:
  - reactants used
  - products formed
  - **ratio** of the **number of moles** of reactants and products i.e. **mole ratio**

- A balanced chemical equation enables us to calculate <u>amount</u> of reactants required and products formed. The amount can be expressed in terms of <u>**mass**</u>, <u>**volume**</u> or <u>**number of moles**</u>.

<u>Example</u>



1 molecule	2 molecules	1 molecule	2 molecules
1 mole	2 moles	1 mole	2 moles

The chemical equation tells us that:

- <u>1</u> molecule of CH_4 reacts with <u>2</u> molecules of O_2 to produce <u>1</u> molecule of CO_2 and <u>2</u> molecules of H_2O.
- <u>1</u> x 6.02 x 10^{23} molecules of CH_4 react with <u>2</u> x 6.02 x 10^{23} molecules of O_2 to produce <u>1</u> x 6.02 x 10^{23} molecules of CO_2 and <u>2</u> x 6.02 x 10^{23} molecules of H_2O.
- <u>1</u> mol of CH_4 reacts with <u>2</u> mol of O_2 to produce <u>1</u> mol of CO_2 and <u>2</u> mol of H_2O.

The relationship between the number of moles of reactants and the number of moles of products involved in a chemical reaction is known as its **stoichiometry.**

<u>Example 1</u>

Calculate the mass of water produced when 0.250 mole of methane is completely burnt in oxygen.

Solution

Step 1: Write the balanced chemical equation.

$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$$

Step 2: Find the **ratio** of the **number of moles** of H_2O to the number of moles of CH_4.

$$\frac{\text{No. of moles of } H_2O}{\text{No. of moles of } CH_4} = \frac{2}{1}$$

Step 3: Use the ratio to find the number of moles of H_2O produced when 0.250 moles of CH_4 is burnt.

No. of mole of H_2O = <u>2</u> x no. of moles of CH_4

= <u>2 x 0.250</u>

= <u>0.500 mol</u>

Step 4: Convert the number of moles of H_2O produced to mass of H_2O.

Mass of H_2O = no. of moles of H_2O x molar mass of H_2O

= <u>0.500 x 18.0</u>

= <u>9.00g (3s.f)</u>

Example 2

Iron(III) oxide reacts with carbon monoxide to form iron and carbon dioxide.
What mass of iron is produced from 40.0 g of iron(III) oxide?

Solution

Step 1: Write the balanced chemical equation.

$Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$

Step 2: Convert 40.0 g of Fe_2O_3 to number of moles of Fe_2O_3.

No. of moles of Fe_2O_3 = $\frac{40.0}{(2 \times 55.8) + (3 \times 16.0)}$ = 0.25063 mol

Step 3: Find the ratio of the number of moles of Fe to the number of moles of Fe_2O_3.

$\frac{\text{No. of moles of Fe}}{\text{No. of moles of } Fe_2O_3} = \frac{2}{1}$

Step 4: Use the ratio to find the number of moles of Fe produced.

No. of moles of Fe produced = $\frac{2}{1}$ x no. of moles of Fe_2O_3 reacted

= $\frac{2}{1} \times 0.25063$

= 0.50126 mol

Step 5: Convert the number of moles of Fe produced to mass of Fe.

Mass of Fe = 0.50126×55.8

= 28.0g (3s.f)

Example 3

Ammonia gas reacts with hot copper(II) oxide to form copper, nitrogen gas and water.

(a) Write a balanced equation for the reaction.

(b) If 2.00 g of copper(II) oxide is used, find the volume of ammonia gas required at r.t.p.

a) $2NH_3 + 3CuO \rightarrow 3Cu + N_2 + 3H_2O$

b) Number of moles of CuO = $\frac{2.00}{(63.5+16.0)}$

= 0.025157 mol

$\frac{\text{No. of moles of } NH_3}{\text{No. of moles of CuO}} = \frac{2}{3}$

No. of moles of NH_3 = $\frac{2}{3} \times 0.025157$

= 0.016771 mol

Volume of NH_3 = 0.016771×24.0

= 0.403 dm^3 (3s.f)

Example 4

Sodium reacts with water to form sodium hydroxide and hydrogen gas. Calculate the mass of sodium that reacts with water to produce 240 cm^3 of hydrogen at r.t.p.

$2Na + 2H_2O \rightarrow 2NaOH + H_2$

No. of moles of $H_2 = \frac{240}{24000}$

$= 0.0100$ mol

$\frac{\text{No. of moles of Na}}{\text{No. of moles of } H_2} = \frac{2}{1}$

No. of moles of Na $= \frac{2}{1} \times$ No. of moles of H_2

$= \frac{2}{1} \times 0.0100$

$= 0.0200$ mol

Mass of Na = no. of moles x molar mass

$= 0.0200 \times 23.0$

$= 0.460$g

Quick Check 1

1 In an experiment, 2.80 g of iron was reacted with excess hydrochloric acid. The equation for the reaction is

$$Fe + 2HCl \rightarrow FeCl_2 + H_2$$

Calculate

(a) the mass of iron(II) chloride produced in the reaction, 6.36

(b) the volume of hydrogen gas produced, measured at r.t.p. 1·20 dm^3

a) No. of moles of Fe $= 2.80 \div 55.8$

$= 0.050179$ mol

$\frac{\text{No. of moles of } FeCl_2}{\text{No. of moles of Fe}} = \frac{1}{1}$

No. of moles of $FeCl_2 = 0.050179$ mol

Mass of $FeCl_2 = 0.050179 \times (55.8 + 35.5 \times 2)$

$= 6.36$g (3s.f)

b) $\frac{\text{No. of moles of } H_2}{\text{No. of moles of Fe}} = \frac{1}{1}$

Volume of $H_2 = 0.050179 \times 24.0$

$= 1.20\ dm^3$ (3s.f)

2 Magnesium burns brilliantly in a gas jar of carbon dioxide to form magnesium oxide and carbon.

(a) Write a balanced equation for the reaction.

(b) If 1.20 g of magnesium is completely burnt, calculate

(i) the volume of carbon dioxide used up, at r.t.p.

(ii) the mass of magnesium oxide formed.

③ Potassium chlorate(V) decomposes when strongly heated. The equation is

$$2KClO_3 \rightarrow 2KCl + 3O_2$$

Calculate the mass of potassium chlorate(V) required to produce 12.0 dm^3 of oxygen gas at r.t.p.

$\frac{\text{No. of moles of } KClO_3}{\text{No. of moles of } 3O_2} = \frac{2}{3}$

No. of moles of O_2 = $3 \times \frac{12.0}{24.0}$

= 1.50 mol

No of moles of $KClO_3$ = $1.50 \times \frac{2}{3}$

= 1

Mass of potassium chlorate required = 2[39.1 + 35.5 + 3(16.0)]

= 245. 2g

6.1 Calculations for Volumes of Reacting Gases

- According to Avogadro's Law, the volume of a gas in a reaction is **proportional** to the number of moles of the gas.
- For reactions involving gaseous reactants and products, the **mole ratio** is **the same** as the **volume ratio** of the gases (provided all volumes are measured at the **same temperature and pressure**).

Example

Nitrogen gas and hydrogen gas reacts to form ammonia.

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

Mole ratio	1	3	2
Volume ratio	1	3	2
Volume /cm^3	10.0	30.0	20.0

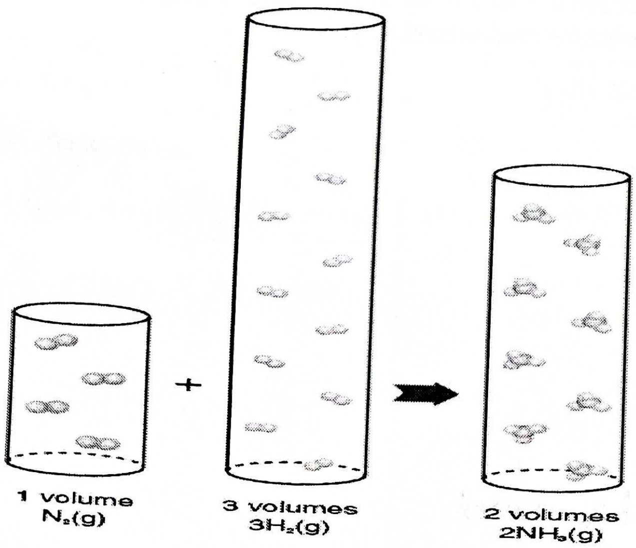
According to the equation,

$$\frac{\text{No. of moles of } H_2}{\text{No. of moles of } N_2} = \frac{3}{1}$$

According to Avogadro's Law,

$$\frac{\text{Volume of } H_2}{\text{Volume of } N_2} = \frac{3}{1}$$

Volume of H_2 = 3 x volume of N_2



We can extend this to the volume of NH_3 produced.

E.g.

If **10.0 cm^3 of N_2** reacts,
the **volume of H_2** that reacts is **30.0 cm^3** and the **volume of NH_3** produced is **20.0 cm^3**.

Example 1
Sulfur dioxide reacts with oxygen according to the equation:

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

40.0 cm^3 of sulfur dioxide reacted with oxygen to form sulfur trioxide.
Calculate
(a) the volume of oxygen reacted.
(b) the volume of sulfur trioxide formed.
(All volumes are measured at the same temperature and pressure.)

Solution
(a) From equation,

$$\frac{\text{No. of moles of } O_2}{\text{No. of moles of } SO_2} = \frac{1}{2}$$

According to Avogadro's Law, $$\frac{\text{volume of } O_2}{\text{volume of } SO_2} = \frac{1}{2}$$

Hence, volume of O_2 reacted = ½ x volume of SO_2 reacted

= ½ x 40.0 cm^3
= 20.0 cm^3

(b) From equation,

$$\frac{\text{No. of moles of } SO_3}{\text{No. of moles of } SO_2} = \frac{1}{1}$$

According to Avogadro's Law, $$\frac{\text{volume of } SO_3}{\text{volume of } SO_2} = \frac{1}{1}$$

Hence, volume of SO_3 formed = $\frac{1}{1}$ x volume of SO_2 reacted
= 20.0 cm^3

Example 2
Ammonia reacts with oxygen according to the equation:

$$4NH_3(g) + 5O_2(g) \rightarrow 4NO(g) + 6H_2O(g)$$

Calculate the volume of ammonia gas that reacted if 100 cm^3 of oxygen gas is used.
(All volumes are measured at the same temperature and pressure.)

From equation,

$$\frac{\text{No. of moles of } NH_3}{\text{No. of moles of } O_2} = \frac{4}{5}$$

According to Avogadro's Law,

$$\frac{\text{Volume of } NH_3}{\text{Volume of } O_2} = \frac{4}{5}$$

Volume of NH_3 reacted = $\frac{4}{5}$ x volume of O_2 reacted
= 80 cm^3

Quick Check 2

1 40.0 cm^3 of ammonia gas was decomposed according to the equation:

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

Calculate

(a) the volume of nitrogen gas produced,

(b) the volume of hydrogen gas produced.

(All volumes are measured at the same temperature and pressure.)

a) From equation,

$$\frac{\text{No. of moles of } N_2}{\text{No. of moles of } NH_3} = \frac{1}{2}$$

According to Avogadro's Law, $\frac{\text{Volume of } N_2}{\text{Volume of } NH_3} = \frac{1}{2}$

Volume of nitrogen gas produced $= 40.0 \times \frac{1}{2}$

$= 20.0 \text{ cm}^3$ //

b) From equation,

$$\frac{\text{No. of moles of } H_2}{\text{No. of moles of } NH_3} = \frac{3}{2}$$

According to Avogadro's Law,

$$\frac{\text{Volume of } H_2}{\text{Volume of } NH_3} = \frac{3}{2}$$

Volume of hydrogen gas produced

$= 40.0 \times \frac{3}{2}$

$= 60.0 \text{ cm}^3$

2 In the motor engine, octane(C_8H_{18}) is burnt according to the equation:

$$2C_8H_{18}(g) + 25O_2(g) \rightarrow 16CO_2(g) + 18H_2O(g)$$

gases

In the hot engine, all the substances in the equation are gases. In the combustion of 40.0 cm^3 of gaseous octane, calculate

(a) the volume of oxygen gas used,

(b) the total volume of gaseous products produced. → Volume of all $\frac{34}{2}$

(All volumes are measured at the same temperature and pressure.)

a) From equation,

$$\frac{\text{No. of moles of } O_2}{\text{No. of moles of } C_8H_{18}} = \frac{25}{2}$$

According to Avogadro's law,

$$\frac{\text{Volume of } O_2}{\text{Volume of } C_8H_{18}} = \frac{25}{2}$$

Volume of O_2 used: $40.0 \times \frac{25}{2}$

$= 500 \text{ cm}^3$ //

b)

$$\frac{\text{No. of moles of gaseous products produced}}{\text{No. of moles of } C_8H_{18}} = \frac{16+18}{2} = \frac{34}{2}$$

According to Avogadro's law

$$= \frac{\text{Volume of gaseous products produced}}{\text{No. of moles of } C_8H_{18}}$$

$= \frac{34}{2}$

Volume of gaseous products produced: $\frac{34}{2} \times 40.0$

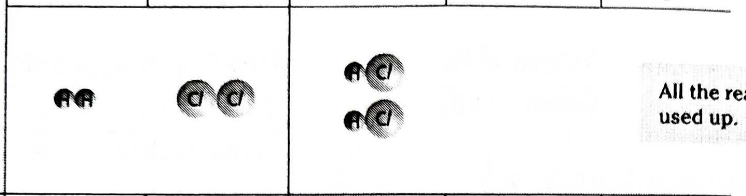
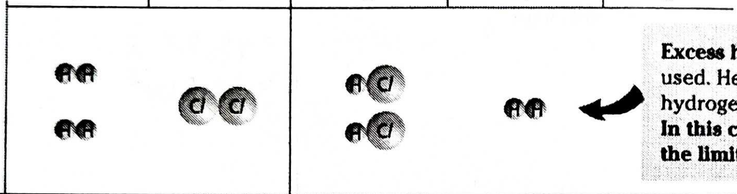
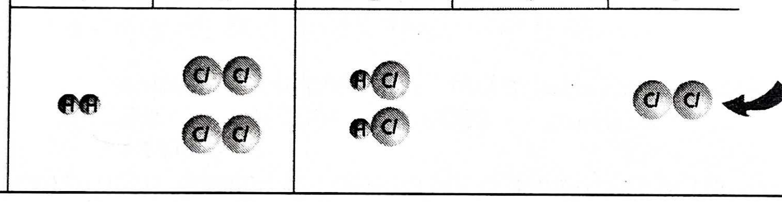
$= 680 \text{ cm}^3$ //

7.1 Limiting Reactant

- The limiting reactant is the reactant that is completely **used up** in a reaction and it **limits** the amount of products formed.

Example: Reaction between hydrogen gas and chlorine gas to give hydrogen chloride gas

$$H_2(g) + Cl_2(g) \rightarrow 2HCl(g)$$

Experiment	Amount of reactants used		Amount of products formed		
	Hydrogen (H_2)	Chlorine (Cl_2)	Hydrogen chloride (HCl)	Unreacted hydrogen	Unreacted chlorine
	1	1	2	0	0
A					
	2	1	2	1	0
B					
	1	2	2	0	1
C					

A: All the reactants are used up.

B: **Excess hydrogen** is used. Hence, some hydrogen is unreacted. **In this case, chlorine is the limiting reactant.**

C: **Excess chlorine** is used. Hence, some chlorine is unreacted. **In this case, hydrogen is the limiting reactant.**

> The amount of a product formed in a reaction is always determined by the amount of the **limiting reactant** (the reactant that is used up).

Example 1

A mixture containing 8.00 g of sulfur and 8.00 g of iron filings is heated to form iron(II) sulfide. Identify the limiting and excess reactant and calculate the mass of iron(II) sulfide formed.

Solution

Step 1: Write the balanced chemical equation. S + Fe = FeS

Step 2: Convert the mass of iron and sulfur given into number of moles.

No. of moles of Fe **given** = 8.00 ÷ 55.8 = 0.14337	No. of moles of S **given** = 8.00 ÷ 32.1 = 0.24922

also includes unreacted sulfur

Take into consideration ratio

Step 3: Find out which reactant is in excess and which reactant is the limiting reactant

$\frac{\text{No. of moles of S}}{\text{No. of moles of Fe}} = \frac{0.24992}{0.14337} = 1.7432$

No. of moles of S **needed** if all the Fe is used up = 1.7432 mol

Since no. of moles of S given > no. of moles of S needed,
S is in **excess** and **Fe** is the **limiting reactant.**

OR

$$\frac{\text{No. of moles of Fe}}{\text{No. of moles of S}} = \frac{0.14337}{0.24992} \cancel{= 0.57366}$$

$\frac{1}{1}$

No. of moles of Fe **needed** if all the S is used up = 0.24992

Since no. of moles of Fe given < no. of moles of Fe needed,

Fe is the **limiting reactant** and **S** is in **excess**.

Step 4: The **number of moles of limiting reactant** determines the number of moles of product.

$$\frac{\text{No. of moles of FeS}}{\text{No. of moles of Fe}} = \frac{1}{1}$$

No. of moles of FeS formed = 0.14337

Step 5: Convert the number of moles of FeS formed into mass.

Mass of FeS produced = 0.14337 × (55.8 + 32.1)

= 12.6g

Example 2

A mixture of 125 g of phosphorus and 323 g of chlorine gas are allowed to react to form phosphorus(III) chloride.

$$P_4(s) + 6Cl_2(g) \rightarrow 4PCl_3(s)$$

Calculate the mass of phosphorus(III) chloride formed.

Solution

No. of moles of P_4 **given** = $\frac{125}{31.0 \times 4}$ = 1.0081mol

No. of moles of Cl_2 **given** = $\frac{323}{35.5 \times 2}$ = 4.5493 mol

$$\frac{\text{No. of moles of } P_4}{\text{No. of moles of } Cl_2} = \frac{1}{6}$$

No. of moles of P_4 **needed** if all the Cl_2 is used up = $\frac{1}{6}$

= $\frac{1}{6}$ × 4.5493 = 0.75822mol

Since no. of moles of P_4 given is > no. of moles of P_4 needed,

P_4 is in excess and Cl_2 is the limiting reactant.

OR

$$\frac{\text{No. of moles of } Cl_2}{\text{No. of moles of } P_4} = \frac{6}{1}$$

No. of moles of Cl_2 **needed** if all the P_4 is used up = $\frac{6}{1} \times 1.0081$

= 6.0486

Since no. of moles of Cl_2 given < no. of moles of Cl_2 needed, Cl_2 is the limiting reactant and P_4 is in excess.

$$\frac{\text{No. of moles of } PCl_3}{\text{No. of moles of } Cl_2} = \frac{4}{6}$$

No. of moles of PCl_3 = $\frac{4}{6} \times 4.5493$

= 3.0329

Mass of PCl_3 = $3.0329 \times [31.0 + 3(35.5)]$

= 3.0329×137.5

= 417 g (3 s.f.)

Example 3 [involving gases]

Propane gas, C_3H_8, burns in air to form carbon dioxide and water.

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

In an experiment, 10.0 cm^3 of propane was burnt in 70.0 cm^3 of oxygen.

(a) Which of the two reacting gases were in excess? Calculate the volume of the excess gas remaining at the end of the experiment.
(b) Calculate the volume of carbon dioxide gas produced.

Solution

(a) From equation,

$$\frac{\text{No. of moles of } O_2}{\text{No. of moles of } C_3H_8} = \frac{5}{1}$$

According to Avogadro's Law, $\frac{\text{volume of } O_2}{\text{volume of } C_3H_8} = \frac{5}{1}$

Volume of O_2 **needed** if all the C_3H_8 is used up = 50.0 cm^3

Since volume of O_2 given < volume of O_2 needed, C_3H_8 is in excess and O_2 is the limiting reactant.

Volume of excess O_2 = $(70.0 + 10.0) - 50.0$

= 30.0 cm^3

(b) $\frac{\text{number of moles of } CO_2}{\text{number of moles of } C_3H_8} = \frac{3}{1}$

$\frac{\text{Volume of } CO_2}{\text{volume of } C_3H_8} = \frac{3}{1}$

Volume of CO_2 = $\frac{3}{1} \times 10.0$

= 30.0 cm^3

Quick Check 3

1 3.60 g of magnesium was burnt in 2.40 dm^3 of oxygen at r.t.p. Calculate the mass of magnesium oxide formed.

$$2Mg + O_2 \rightarrow 2MgO$$

From equation,

$$\frac{\text{No. of moles of } O_2}{\text{No. of moles of Mg}} = \frac{1}{2}$$

According to Avogadro's Law, $\frac{\text{Volume of } O_2}{\text{Volume of Mg}} = \frac{1}{2}$

Volume of Mg $= 2.40\,dm^3 \times 2$

$= 4.80\,dm^3$

No. of moles of Mg needed to react. $= \frac{4.80}{24}$

$= 0.200\,mol$

2 15.0 cm^3 of ethene (C_2H_4) was burnt in 80.0 cm^3 of oxygen. What is the volume of the resulting gas mixture if all volumes are measured at r.t.p.?

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

8.1 Concentration of Solutions

- The concentration of a solution tells us the <u>amount</u> of <u>solute</u> in a unit volume of a solution. There are two ways of measuring the concentration:

1. The **mass (in grams) of solute in 1 dm³ of <u>solution</u>** (g/ dm³)
   A concentration of 10.0 g/dm³ means there are <u>10.0 g</u> of solute in every <u>1</u> dm³ of solution.

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

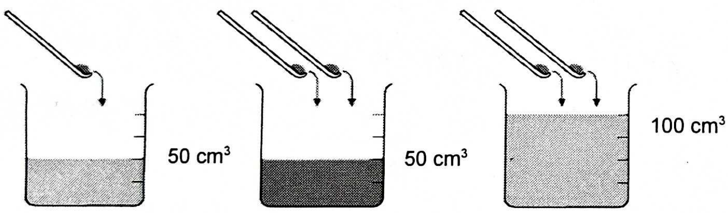
2. The **moles of solute in 1 dm³ of <u>solution</u>** (mol/ dm³) is known as the **<u>molar</u> concentration or molarity.**
   A sodium hydroxide solution with a concentration of 0.500 mol/dm³ has <u>0.500</u> moles of NaOH dissolved in every <u>1</u> dm³ of solution.

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

- A convenient formula involving molar concentration and concentration in g/dm³

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

Comparing concentrations



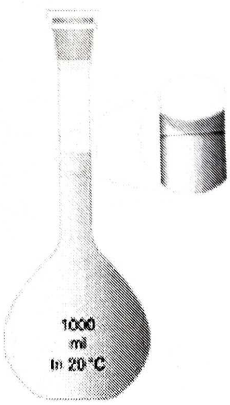
Solution A **Solution B** **Solution C**

➢ Solution A and B have the same <u>volume</u>, but Solution <u>B</u> has **twice** the amount of <u>solute</u> as Solution <u>A</u> and thus has twice the concentration of Solution <u>A</u>.

➢ Solution B and C have the same amount of <u>solute</u>, but Solution <u>C</u> has **twice** the volume as Solution <u>B</u> and thus has half the concentration of Solution <u>B</u>.

Standard Solution

- A **standard solution** is a solution of **known concentration.**
- It is usually made by dissolving a known mass of solute in a known volume of solution in a volumetric flask.



Example 1
A 200 cm^3 solution contains 10.0 g of potassium hydroxide, KOH. Calculate the concentration of potassium hydroxide in (a) g/dm^3, and (b) mol/dm^3.

Solution

(a) Concentration of KOH solution $= \dfrac{\text{mass of solute}}{\text{volume of solution}}$

$= 10.0 \div \frac{200}{100}$

$= 50.0 g/dm^3$

(b) Molar mass of KOH = 39.1 + 16.0 + 1.0 = 56.1 g/mol

Concentration of KOH solution $= \dfrac{\text{concentration in } g/dm^3}{\text{molar mass}}$

$= \cancel{56.1} \quad \frac{50.0}{56.1}$

$= 0.891 mol/dm^3$

(c) Alternative answer

Number of moles of KOH $= \dfrac{\text{mass}}{\text{molar mass}}$

$= \dfrac{10.0}{39.1+16.0+1.0} = 0.17825 mol$ (5 s.f)

Concentration of KOH solution $= \dfrac{\text{number of moles of solute}}{\text{volume of solution}}$

$= \dfrac{0.17825}{0.2}$

$= 0.891 mol/$ ~~dm~~ dm^3 (3 s.f)

Example 2
What mass of H_2SO_4 must be dissolved in 250 cm^3 of solution to prepare a 0.200 mol/dm^3 solution?
↳ 0.25

No. of moles of H_2SO_4 = concentration (mol/dm^3) x volume

$= 0.200 \times 0.25$

$= 0.05 mol$

Mass of H_2SO_4 = no. of moles x molar mass

$= 0.05 \times [2(1.0) + 32.1 + 4(16.0)]$

$= 4.905 g$ //

Quick check 4
1 Calculate the concentration of the following solutions, in mol/dm^3:
(a) 0.200 mol of potassium hydroxide in 4.00 dm^3 of solution.
(b) 0.0100 mol of sulfuric acid in 50.0 cm^3 of solution.

a) Concentration of KOH solution $= \dfrac{0.200}{4.00}$

$= 0.0500 mol/dm^3$ //

b) Concentration of H_2SO_4 solution $= \left(\frac{50.0}{1000}\right) \div 0.05$

$= 0.200 mol/dm^3$ //

2 Calculate the number of moles of
(a) NaOH in 25.0 cm^3 of 0.400 mol/dm^3 NaOH
(b) NaC*l* in 500 cm^3 of 0.0200 mol/dm^3 NaC*l*
(c) H_2SO_4 in 2.00 dm^3 of 4.90 g/dm^3 of H_2SO_4

a) Number of moles of NaOH

$= \frac{25.0}{1000} \times 0.400$

$= 0.0100 mol$

b) Number of moles of NaCl

$= \frac{500}{1000} \times 0.0200$

$= 0.0100 mol$

c) Molar concentration of H_2SO_4

$= \frac{4.90}{(2\times1.0+32.1+4\times16.0)}$

$= 0.049949 mol/dm^3$

No. of moles of H_2SO_4
= concentration × volume
$= 0.049949 \times 2$
$= 0.0999 mol$

3 Calculate the mass of
(a) HC*l* in 200 cm^3 of 1.00 mol/dm^3 hydrochloric acid.
(b) HNO_3 in 1000 cm^3 of 0.100 mol/dm^3 nitric acid.

a) Molar mass of HCl = 36.5
g/dm^3 of HCl = 36.5×100
$= 36.5 g/dm^3$
$200 cm^3 = 0.21 dm^3$
Mass of HCl $= 36.5 \times 0.2$
$= 7.3 g$

b) $1000 cm^3 = 1 dm^3$
Molar mass of HNO_3
$= 63$
g/dm^3 of $HNO_3 = 63 \times 0.600$
$= 6.3$
Mass of $HNO_3 = 6.3 \times 1$
$= 6.30 g$

4 What mass of potassium hydroxide must be dissolved in 100 cm^3 of solution to prepare a 0.400 mol/dm^3 solution?

$0.400 mol/dm^3$
$100 cm^3 = 0.1 dm^3$
Number of moles of KOH required $= 0.400 \times (0.1)$
$= \cancel{0.40}\ 0.0400 mol$

Mass of KOH $= 0.04 \times 56.105$
$= 2.2440 g$
$= 2.24 g$ (3 s.f)

9 Percentage Yield and Percentage Purity

9.1 Percentage Yield

(d) The amount of product formed in a reaction is known as the yield.
(e) The **theoretical yield** is the amount of product **calculated from the equation**.
(f) The **experimental yield** is the amount of product that is **actually obtained in the experiment**.
(g) In any reaction, the experimental yield is almost always **lower** than the theoretical yield.

$$\textbf{Percentage yield} = \frac{\textbf{Experimental yield}}{\textbf{Theoretical yield}} \times 100\ \%$$

Example 1

128 g of sulfur dioxide, SO_2, was reacted with oxygen to produce sulfur trioxide, SO_3.
The equation for the reaction is:

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

140 g of SO_3 was produced in the reaction. Calculate the percentage yield of the SO_3.

Solution

No. of moles of SO_2 = $\frac{128}{32.1+2(16.0)}$ = 1.9969 (5s.f)

From the equation, $\frac{\text{no. of moles of } SO_3}{\text{no. of moles of } SO_2} = \frac{2}{2} = \frac{1}{1}$

No. of moles of SO_3 = 1 x 1.9969
= 1.9969 mol

Theoretical mass of SO_3 produced = no. of moles x molar mass
= 1.9969 x (32.1 + 3 x 16.0)
= 159.95 g

% yield of SO_3 = $\frac{\text{Experimental yield of } SO_3}{\text{Theoretical yield of } SO_3} \times 100\ \%$
= $\frac{140}{159.95} \times 100\%$
= 87.5%

9.2 Percentage Purity

- Sometimes, substances are not pure and the percentage purity needs to be calculated.

$$\textbf{Percentage purity} = \frac{\textbf{mass of pure substance}}{\textbf{mass of impure sample}} \times 100\ \%$$

Example 1

An impure sample of calcium carbonate ($CaCO_3$) contains calcium sulfate as an impurity. When excess hydrochloric acid was added to 6.00 g of the sample, 1200 cm³ of gas were produced (measured at r.t.p.).

Calculate the percentage purity of the calcium carbonate sample.

Solution

Equation : $CaCO_3 + 2HCl \rightarrow CaCl_2 + H_2O + CO_2$

(Note: Only $CaCO_3$ will react with hydrochloric acid but not the $CaSO_4$ impurity)

Question : *Why is hydrochloric acid added in excess?* To ensure all the $CaCO_3$ is reacted

No. of moles of CO_2 produced = $\frac{\text{volume}}{\text{molar volume}}$

$= \frac{1200}{24000} = 0.0500 \text{ mol}$

From equation, $\frac{\text{no. of moles of } CaCO_3}{\text{no. of moles of } CO_2} = \frac{1}{1}$

no. of moles of $CaCO_3$ reacted = $\frac{1}{1} \times 0.0500 \text{ mol}$

$= 0.0500 \text{ mol}$

Mass of $CaCO_3$ reacted = no. of moles x molar mass

$= 0.0500 \times (40.1 + 12.0 + 3 \times 16.0)$

$= 5.005 g$

% purity of $CaCO_3$ in sample = $\frac{\text{mass of pure } CaCO_3}{\text{mass of impure sample}} \times 100\ \%$

$= \frac{5.005}{6.00} \times 100\%$

$= 83.4\%$

Quick Check 5

1 In the oxidation of 1 mole of ethanol (C_2H_6O), 1 mole of ethanoic acid ($C_2H_4O_2$) is obtained. If 9.20 g of ethanol is used and the mass of ethanoic acid obtained is 9.00 g, calculate the percentage yield of ethanoic acid.

$\frac{\text{Number of moles of } C_2H_6O}{\text{Number of moles of } C_2H_4O_2} = \frac{1}{1}$

No. of moles of $C_2H_4O_2$ produced = $\frac{\text{mass}}{\text{molar mass}}$

$= \frac{9.00}{2(12.0)+4(1.0)+2(16.0)}$

$= 0.1500 \text{mol}$

Expected yield of ethanoic acid

$= \frac{9.20}{2(12.0)+6(1.0)+(16.0)}$

$= 0.2000 \text{mol}$

Percentage yield of ethanoic acid

~~$\frac{0.1500}{0.2000}$~~ $\frac{0.1500}{0.2000} \times 100\%$

$= 75\%$

2 0.500 g of impure sodium carbonate is dissolved in excess dilute hydrochloric acid. The volume of carbon dioxide collected, measured at r.t.p., is found to be 105 cm^3.
What is the percentage purity of the sodium carbonate?

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

No. of moles of CO_2 produced $= \frac{\text{Volume}}{\text{molar volume}}$

$= \frac{0.105}{24.0}$

$= 0.0043750$ mol (5 s.f.)

10.1 Volumetric Analysis - Calculations for Titration

Titration is a method of volumetric analysis. It is used to **determine the concentrations of solutions.**
Titration involves measuring volumes of solutions and calculations using
- a solution of **unknown** concentration;
- a solution of **known** concentration.

Refer to Acid, Bases and Salts chapter on salt preparation using titration method.

Example 1
30.0 cm^3 of 0.100 mol/dm^3 sodium hydroxide solution reacted completely with 25.0 cm^3 of sulfuric acid in a titration. Calculate the concentration of sulfuric acid in
(a) mol/dm^3,
(b) g/dm^3.

Solution
Step 1: Write the balanced chemical equation.

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

Step 2: Find the number of moles of NaOH

30.0 cm^3 = 0.03

No. of moles of NaOH = concentration (mol/dm^3) x volume in dm^3

= 0.100 × 0.0300

= 0.00300 mol

Step 3: Find the ratio of the number of moles of H_2SO_4 to the number of moles of NaOH.

$$\frac{\text{No. of moles of } H_2SO_4}{\text{No. of moles of NaOH}} = \frac{1}{2}$$

Step 4: Use the ratio to find the number of moles of H_2SO_4 produced.

No. of moles of H_2SO_4 = $\frac{1}{2}$ x no. of moles of NaOH

= $\frac{1}{2}$ x 0.00300

= 0.00150 mol

Step 5: Find the concentration of H_2SO_4 in mol/dm³.

Concentration of H_2SO_4 = no. of moles of H_2SO_4 ÷ volume in dm³

= 0.00150 ÷ 0.03

= 0.05 mol/dm³

Step 6: Find the concentration of H_2SO_4 in g/dm³.

Concentration of H_2SO_4 (g/dm³) = concentration of H_2SO_4 (mol/dm³) x molar mass

= 0.05 mol/dm³ x (2 x 1.0 + 32.1 + 4 x 16.0)

= 5.89 g/dm³

Example 2

In a titration, it was found that 15.0 cm³ of a solution of hydrochloric acid will exactly neutralise 25.0 cm³ of a 1.20 g/dm³ solution of sodium hydroxide. Find the concentration of the hydrochloric acid solution in mol/dm³.

Solution

Equation: $HCl + NaOH \rightarrow H_2O + NaCl$

Concentration of NaOH (mol/dm³) = concentration (g/dm³) ÷ molar mass

= ~~10~~ 1.20 g/dm³ ÷ (23.0 + 16.0 + 1.0)

= 0.0300 mol/dm³

No. of moles of NaOH = concentration (mol/dm³) x volume (dm³) 0.025 dm³

= 0.0300 mol/dm³ x 0.025 dm³

= 0.00075 mol

$$\frac{\text{No. of moles of HCl}}{\text{No. of moles of NaOH}} = \frac{1}{1}$$

No. of moles of HC*l* = no. of moles of NaOH
= 0.00075 mol

Concentration of HC*l* = no. of moles of HC*l* ÷ volume in dm^3

= 0.00075 mol ÷ 0.015

= 0.5 mol/dm³

Example 3
2.80 g of an acid, H_3X, were dissolved in water and made up to 250 cm^3. A 25.0 cm^3 portion of this solution was neutralised by 28.6 cm^3 of 0.300 mol/dm^3 sodium hydroxide solution according to the equation:

$$H_3X\ (aq) + 3NaOH\ (aq) \rightarrow Na_3X\ (aq) + 3H_2O\ (l)$$

Calculate
(a) the mass of the acid in 1 dm^3 of solution,
(b) the number of moles of the acid in 1 dm^3 of solution,
(c) the relative molecular mass of the acid.

Solution
(a) Mass of acid in 1 dm^3 = mass ÷ volume in dm^3

= 2.80 ÷ 0.25

= 11.2 g //

(b) No. of moles of NaOH = concentration(mol/dm^3) × volume in dm^3

= 0.0300 × 0.0286

= 0.00858 mol

$$\frac{\text{No. of moles of } H_3X}{\text{No. of moles of NaOH}} = \frac{1}{3}$$

No. of moles of H_3X = $\frac{1}{3}$ × 0.00858

= 0.000286 mol

No. of moles of H_3X in 1 dm^3 = no. of moles ÷ volume in dm^3

= 0.000286 ÷ 0.025

= 10.1144 mol (4s.f)

= 0.114 mol // (3s.f)

(c) Molar mass of H_3X = concentration(g/dm^3) ÷ concentration(mol/dm^3)

= 11.2 ÷ 0.1144

= 97.9 g/mol

Relative molecular mass of H_3X = 97.9

Example 4

16.0 g of **impure** sodium carbonate are dissolved in 1 dm^3 of solution. A 25.0 cm^3 portion of this solution required 30.0 cm^3 of 0.200 mol/dm^3 hydrochloric acid for reaction in a titration. Calculate the percentage purity of the sodium carbonate.

Solution

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

No. of moles of HC*l* = concentration (mol/dm^3) × volume in dm^3 = 0.200 × 0.0300

= 0.00600 mol

$$\frac{\text{No. of moles of } Na_2CO_3}{\text{No. of moles of HC}l} = \frac{1}{2}$$

No. of moles of pure Na_2CO_3 in 25.0 cm^3 of solution = $\frac{1}{2}$ × 0.00600

= 0.00300 mol

No. of moles of pure Na_2CO_3 in 1 dm^3 of solution = no. of moles ÷ volume in dm^3

= $\frac{0.00300}{0.0250}$

= 0.120 mol

Mass of pure Na_2CO_3 in 1 dm^3 = no. of moles in 1 dm^3 × molar mass

= 0.120 × (2 × 23.0) + (12.0) + (16 × 0.3)

= 0.120 × 106

= 12.72 g

% purity of Na_2CO_3 in sample = $\frac{\text{mass of pure } Na_2CO_3}{\text{mass of impure } Na_2CO_3}$ x 100 %

= $\frac{12.72}{16.0}$ × 100%

= 79.5%

Quick Check 6 0.032 dm^3

1 32.0 cm^3 of 0.100 mol/dm^3 NaOH reacts with 25.0 cm^3 of HC*l* in a titration. Calculate the concentration of the HC*l* in (a) mol/dm^3, and (b) in g/dm^3.

NaOH + HCl = H_2O + ~~H~~ NaCl

Number of moles of NaOH = 0.032 × 0.100

= 0.0032 mol

$\frac{\text{No. of moles of HCl}}{\text{No. of moles of NaOH}} = \frac{1}{1}$

∴ No. of moles of HCl = 0.0032 mol

a) Concentration of HCl in mol/dm^3 = 0.0032 ÷ ($\frac{25.0}{1000}$)

= 0.128 mol/dm^3

b) Concentration of HCl in ~~molar~~ ~~mass~~ g/dm^3

= 0.128 × (1.0 + 35.5)

= 4.672 ~~mo~~ g/dm^3

2 A solution of NaOH contains 6.00 g/dm³ of NaOH. 20.0 cm³ of the NaOH reacts completely with 40.0 cm³ of HNO_3. Calculate

(a) the concentration of NaOH in mol/dm³.

(b) the number of moles of HNO_3 in 1 dm³.

(c) the mass of HNO_3 in 1 dm³.

$NaOH + HNO_3 \rightarrow NaNO_3 + H_2O$

Mass of NaOH: $\frac{20.0}{1000} \times 6.00 g/dm^3$

$= 0.12g$

3 7.40 g of a sample of impure potassium carbonate was dissolved in water and made up to 1 dm³. 20.0 cm³ of this solution required 19.80 cm³ of 0.0500 mol/dm³ sulfuric acid for neutralisation. What is the percentage purity of the potassium carbonate?

4 In a titration experiment, 22.5 cm³ of 7.00 g/dm³ nitric acid neutralised 20.0 cm³ of an alkali, XOH, containing 5.00 g of alkali per dm³. Calculate the relative atomic mass of X.

ANSWERS TO QUICK CHECK

	Quick Check 1 (pg 21 - 22)
1(a)	6.36 g
1(b)	1.20 dm^3
2(a)	$2Mg + CO_2 \rightarrow 2MgO + C$
2(b)(i)	0.593 dm^3
2(b)(ii)	1.99 g
3	40.9 g
	Quick Check 2 (pg 24)
1(a)	20.0 cm^3
1(b)	60.0 cm^3
2(a)	500 cm^3
2(b)	680 cm^3
	Quick Check 3 (pg 28)
1	5.97 g
2	65.0 cm^3

	Quick Check 4 (pg 30 - 31)
1(a)	0.0500 mol/dm^3
1(b)	0.200 mol/dm^3
2(a)	0.0100 mol
2(b)	0.0100 mol
2(c)	0.0999 mol
3(a)	7.30 g
3(b)	6.30 g
4	2.24 g
	Quick Check 5 (pg 33 - 34)
1	75.0 %
2	92.8 %
	Quick Check 6 (pg 37 - 38)
1(a)	0.128 mol/dm^3
1(b)	4.67 g/dm^3
2(a)	0.150 mol/dm^3
2(b)	0.0750 mol
2(c)	4.73g
3	92.4 %
4	23.0

Mode Concept Summary Exercise

72.0 g of $C_6H_{12}O_6$

Mass of oxygen in the sample

$\frac{6(16.0)}{180.0} \times 72.0$

$= 38.4g$

molar mass has units unlike M_r

Molar mass of the compound

$6(12.0)+12(1.0)+6(16.0)$

$= 180 g/mol$

Percentage of carbon in the sample

$\frac{6(12.0)}{180.0} \times 100\% = 40.0\%$

Molarity of the solution (mol/dm³)

$\frac{0.400}{250 \times 10^{-3}}$

$= 1.60 mol/dm^3$

Number of molecules in the sample

$0.400 \times 6.02 \times 10^{23}$

$= 2.408 \times 10^{23}$

$= 2.41 \times 10^{23}$ molecules

Number of moles of molecules of the compound

$\frac{72.0}{6(12.0)+12(1.0)+6(16.0)}$

$= 0.400 mol$

Dissolve in 250 cm³ of water to form solution

Number of atoms in the sample

$0.400 \times 6.02 \times 10^{23} \times 24$

$= 5.78 \times 10^{24}$

Concentration of the solution in g/dm³

1.60×180.0

$= 288 g/dm^3$

complete combustion with oxygen to form carbon dioxide and water only

Equation: $C_6H_{12}O_6 + 6O_2 \rightarrow 6CO_2 + 6H_2O$

Volume of carbon dioxide produced (at r.t.p.)

No. of moles of CO_2 produced $= 6 \times 0.400 = 2.40 mol$

Volume of CO_2 produced

$= 2.40 \times 24.0 = 57.6 dm^3$

Volume of air needed assuming air contains 20% oxygen (at r.t.p.)

Moles of O_2 needed $= 6 \times 0.400 \times 2.40$

Volume of $O_2 = 2.40 \times 24.0 = 57.6$

Volume of air $= \frac{57.6}{20} \times 100$

$= 288 dm^3$

Mass of water produced

No of moles of H_2O produced $= 6 \times 0.400 = 2.40 mol$

Mass of H_2O produced $= 2.40 \times 18.0 = 43.2g$