Raffles Institution
Year 5 H2 Chemistry 2025
Supplementary Notes – The Mole Concept and Stoichiometry

Content

- Relative masses of atoms and molecules
- The mole, the Avogadro constant
- The calculation of empirical and molecular formulae
- Reacting masses and volumes (of solutions and gases)

Learning Outcomes

[The term relative formula mass or M_r will be used for ionic compounds]

Candidates should be able to:

(a) define the terms relative *atomic*, *isotopic*, *molecular* and *formula masses*
(b) define the term *mole* in terms of the Avogadro constant
(c) calculate the relative atomic mass of an element given the relative abundances of its isotopes
(d) define the terms *empirical* and *molecular formula*
(e) calculate empirical and molecular formulae, using combustion data or composition by mass
(f) write and/or construct balanced equations
(g) perform calculations, including use of the mole concept involving:
  (i) reacting masses (from formulae and equations);
  (ii) volumes of gases (e.g. in the burning of hydrocarbons);
  (iii) volumes and concentrations of solutions.
  [when performing calculations, candidates' answers should reflect the number of significant figures given or asked for in the question]
(h) deduce stoichiometric relationships from above calculations such as those in (g)

Lecture Outline

1 Atoms and Sub-atomic Particles
2 Relative Masses
3 The Mole and Related Concepts
4 Empirical and Molecular Formulae
5 Stoichiometry
6 Reacting Volumes of Gases
7 Concentration of a Solution
8 Acid-Base Titrations

References

1 Cambridge International AS and A Level Chemistry (by Peter Cann and Peter Hughes)
2 Cambridge International AS and A Level Chemistry Coursebook with CD-ROM (by Lawrie Ryan and Roger Norris)
3 A Level Chemistry 4ed (by E. N. Ramsden)
4 Chemistry in Context 6ed (by Graham Hill and John Holman)
5 Chemistry – The Molecular Nature of Mater and Change (by Silberberg)
6 https://www.chemguide.co.uk

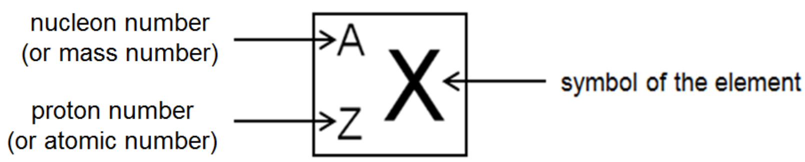
1 Atoms and Sub-atomic Particles

1.1 The sub-atomic particles

- An atom is the smallest part of an element which can ever exist, whereas a molecule is the smallest part of an element or a compound which can exist alone under ordinary conditions.
- All atoms are composed of three fundamental particles — protons, neutrons and electrons.

Sub-atomic particle	**proton**	**neutron**	**electron**
Symbol	${}^{1}_{1}p$	${}^{1}_{0}n$	${}^{0}_{-1}e$
Relative mass	1	1	$\frac{1}{1840}$
Relative charge	+1	0	–1
Location within the atom	in the nucleus	in the nucleus	around the nucleus

1.2 Important terms and definitions

	Term	**Symbol**	**Definition**
(a)	proton number or atomic number	Z	• The proton number (or atomic number) of an element is the number of protons in the nucleus of an atom of that element. • The atomic number determines the identity of an atom. For example, every atom with an atomic number of 6 is a carbon atom; it contains 6 protons in its nucleus.
(b)	nucleon number or mass number	A	• The nucleon number (or mass number) of an element is the total number of protons and neutrons in the nucleus of an atom of that element. • **Note:** Protons and neutrons are collectively known as nucleons because they are both found in the nucleus.
(c)	nuclide	${}^{A}_{Z}X$	• A nuclide is any species of given mass number and atomic number. • Examples: ${}^{1}_{1}H$ ${}^{9}_{4}Be$ ${}^{12}_{6}C$ ${}^{16}_{8}O$ • The nuclide of an element is represented by  • **Note:** Total number of protons and neutrons = A Number of protons = Z Number of neutrons = A – Z Number of electrons (for uncharged species) = number of protons = Z

1.3 Isotopes

- Isotopes of an element are atoms with the same proton number but different nucleon numbers (i.e. they have the same number of protons but different number of neutrons in the nucleus).
- Isotopes have the same number of electrons ⇒ the same chemical properties
  Isotopes have different numbers of neutrons (i.e. different masses) ⇒ different physical properties
- Most elements consist of mixtures of isotopes. The abundance of each isotope in the mixture is called its isotopic abundance (in terms of percentages or fractions).
- Example 1: Isotopes of hydrogen

Name	Symbol	Number of protons	Number of neutrons	Number of electrons	Isotopic abundance in natural hydrogen
protium (hydrogen)	${}^{1}_{1}H$ or H	1	0	1	99.984%
deuterium (heavy hydrogen)	${}^{2}_{1}H$ or D	1	1	1	0.015%
tritium	${}^{3}_{1}H$ or T	1	2	1	very rare — 1 part in 10^{17} (radioactive and unstable)

- Example 2: Isotopes of chlorine

Name	Symbol	Number of protons	Number of neutrons	Number of electrons	Isotopic abundance in naturally occurring chlorine
chlorine-35	${}^{35}_{17}Cl$	17	18	17	75%
chlorine-37	${}^{37}_{17}Cl$	17	20	17	25%

Worked Example 1

(a) Which one of the following particles has more electrons than protons and more protons than neutrons?

A D^- **B** OH^- **C** H_3O^+ **D** OD^-

(b) Consider the atoms and ions given in the table below.

Atom / Ion	Number of electrons	Number of neutrons
P	11	15
Q^{2-}	11	17
R^+	10	15
S^-	12	17
T^+	13	16

Which of the following is an isotope of P?

A Q **C** S
B R **D** T

Solution

(a)

	D^-	OH^-	H_3O^+	OD^-
No. of electrons	2	10		10
No. of protons	1	9		9
No. of neutrons	1	8		9

(b)

	P	Q^{2-}	R^+	S^-	T^+
No. of electrons	11	11	10	12	13
No. of protons					
No. of neutrons	15	17	15	17	16

2	Relative Masses

2.1 The carbon-12 scale

- The masses of atoms are very small, from 10^{-24} to 10^{-22} grams. Chemists use a relative atomic mass scale to compare the masses of different atoms.
- In 1961, the carbon-12 atom was adopted by the International Union of Pure and Applied Chemistry (IUPAC) as the reference standard for relative atomic masses.
- On the carbon-12 scale, atoms of the isotope ^{12}C are assigned a relative atomic mass of 12 and the relative masses of all other atoms are obtained by comparison with the mass of the carbon-12 atom.

2.2 Relative isotopic mass

- The relative isotopic mass of a particular isotope of an element is defined as follows:

$\text{Relative isotopic mass} = \dfrac{\text{mass of 1 atom of the isotope}}{\frac{1}{12} \times \text{mass of 1 atom of carbon-12}}$	**Note:** No units

- Examples: Relative isotopic mass of ^{21}Ne = 20.994 ≈ 21.0
  Relative isotopic mass of ^{35}Cl = 34.97 ≈ 35.0
- Why is the relative isotopic mass of an isotope very close to a whole number?

  Reason: On the carbon-12 scale, the relative masses of the proton and neutron are both very close to one and the electron has a negligible mass. It therefore follows that all relative isotopic masses will be very close to whole numbers which are essentially the corresponding nucleon numbers of the isotopes.
- In calculations, the relative isotopic mass of an isotope is often approximated by the nucleon number of that isotope. The two are assumed to be identical in all but the most accurate work.

2.3 Relative atomic mass (Symbol: A_r)

- The relative atomic mass (A_r) of an element is defined as follows:

$\text{Relative atomic mass} = \dfrac{\text{(weighted) average mass of 1 atom of the element}}{\frac{1}{12} \times \text{mass of 1 atom of carbon-12}}$	**Note:** No units

- The A_r values can be found in the Periodic Table given in the Data Booklet.
  Example: A_r of chlorine = 35.5

13	14	15	16	17	18
Al	Si	P	S	Cl	Ar
aluminium	silicon	phosphorus	sulfur	chlorine	argon
27.0	28.1	31.0	32.1	35.5	39.9

- The relative atomic mass of an element may not be close to a whole number. Why is this so?
  Reason: This is because naturally occurring elements often consist of a mixture of isotopes and the relative atomic mass of an element takes into consideration the different isotopes and their relative abundances.

Worked Example 2

Calculate the relative atomic mass of chlorine from the given data.

Isotope	Relative isotopic mass	Percentage abundance
^{35}Cl	34.97	75.53
^{37}Cl	36.95	24.47

Solution

A_r of Cl = $\dfrac{(75.53)(34.97) + (24.47)(36.95)}{75.53 + 24.47}$ = **35.5** (3 s.f.)

Worked Example 3

The isotopes ^{79}Br and ^{81}Br have accurate isotopic masses of 78.918 and 80.916 respectively. Given that the A_r of bromine is 79.904, calculate the percentage abundances of these two isotopes.

Solution

Let x be the percentage abundance of ^{79}Br and (100 – x) be the percentage abundance of ^{81}Br.

A_r of Br =

Percentage abundance of ^{79}Br = x =

Percentage abundance of ^{81}Br = 100 – x =

2.4 Relative molecular mass (Symbol: M_r)

- This is the term used when referring to the relative masses of molecular elements or covalent compounds.

Relative molecular mass = $\dfrac{\text{(weighted) average mass of 1 molecule of the substance}}{\frac{1}{12} \times \text{ mass of 1 atom of carbon-12}}$	**Note:** No units

- M_r of a substance = sum of the A_r of all the constituent atoms shown in the molecular formula
- Examples: M_r of O_2 = (2)(16.0) = 32.0

  M_r of aspirin, $C_9H_8O_4$ = (9)(12.0) + (8)(1.0) + (4)(16.0) = 180.0

2.5 Relative formula mass (Symbol: M_r)

- This is the term used when referring to the relative masses of ions or ionic compounds.

Relative formula mass = $\dfrac{\text{(weighted) average mass of 1 formula unit of the substance}}{\frac{1}{12} \times \text{ mass of 1 atom of carbon-12}}$	**Note:** No units

- **Note:** A formula unit is the smallest collection of atoms from which the formula of a compound can be established.
- Examples: Relative formula mass of NaCl = 23.0 + 35.5 = 58.5

  M_r of $CuSO_4.5H_2O$ = 63.5 + 32.1 + (4)(16.0) + (10)(1.0) + (5)(16.0) = 249.6

3 The Mole and Related Concepts

3.1 The mole and the Avogadro constant

- The mole is the SI (Système International) unit for measuring "amount of substance".

Base physical quantity	Usual symbol	SI unit	Symbol for unit
amount of substance	n	mole	mol

- Experimentally, it has been found that 12 grams of carbon-12 contain 6.02 x 10^{23} carbon atoms.
  $\Rightarrow$ 1 mole of carbon-12 contains 6.02 x 10^{23} carbon-12 atoms.
- The definition of the mole reads as follows:

> A mole of substance is the amount of that substance which contains 6.02 x 10^{23} elementary entities of that substance.
>
> **Note:** The elementary entities must be specified and may be atoms, molecules, ions, electrons, other particles, or other formula units. The quantity of 6.02 x 10^{23} mol^{-1} is termed the Avogadro constant (Symbol: *L*)
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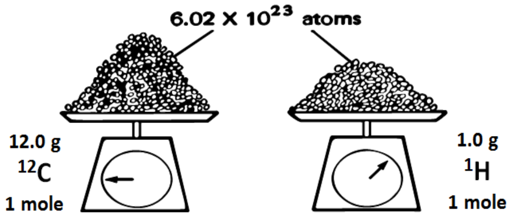
> 
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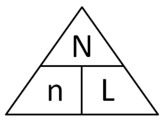
- The Avogadro constant (*L)* is the constant of proportionality between the number of specified entities of a substance (*N*) and the amount of specified entities of that substance (n).

number of specified entities of a substance	=	Avogadro constant (mol^{-1})	x	amount of specified entities of that substance (mol)
N	=	*L*	x	n



- Relationship between the mole and Avogadro constant

> **1 mole** of H_2O
> - contains **6.02 x 10^{23}** H_2O molecules.
> - contains **1 mole** of O atoms and hence contains **6.02 x 10^{23}** O atoms.
> - contains **2 moles** of H atoms and hence contains **(6.02 x 10^{23})(2)** H atoms.
> - contains **3 moles** of atoms and hence contains **(6.02 x 10^{23})(3)** atoms.
>
> **1 mole** of $MgCl_2$
> - contains **6.02 x 10^{23}** formula units of $MgCl_2$.
> - contains **1 mole** of Mg^{2+} ions and hence contains **6.02 x 10^{23}** Mg^{2+} ions.
> - contains **2 moles** of Cl^- ions and hence contains **(6.02 x 10^{23})(2)** Cl^- ions.
> - contains **3 moles** of ions and hence contains **(6.02 x 10^{23})(3)** ions.

3.2 <u>Molar mass</u>

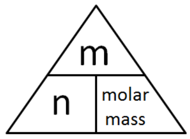
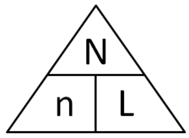
- The <u>molar mass of a substance</u> is the mass of one mole of the substance. It has units of g mol^{-1}.
- Examples:

Molar mass of Fe	= 55.8 g mol^{-1}	A_r of Fe	= 55.8
Molar mass of H_2O	= 18.0 g mol^{-1}	M_r of H_2O	= 18.0
Molar mass of $MgCl_2$	= 95.3 g mol^{-1}	M_r of $MgCl_2$	= 95.3
Molar mass of OH^-	= 17.0 g mol^{-1}	M_r of OH^-	= 17.0

- Note that the molar mass of a substance has the same numerical value as the A_r or M_r of that substance except that it has units of g mol^{-1} while both A_r and M_r have no units.

3.3 <u>Relationship between amount, mass, molar mass and number of particles</u>

- Take note of the following:

Amount of substance X (mol) $= \dfrac{\text{mass of X (g)}}{\text{molar mass of X (g mol}^{-1})}$	
or n_X (mol) $= \dfrac{m_X \text{ (g)}}{\text{molar mass of X (g mol}^{-1})}$	m_X = mass of substance X / g n_X = amount of X / mol
Amount of substance X (mol) $= \dfrac{\text{number of particles of X}}{\text{Avogadro constant (mol}^{-1})}$	
or n_X (mol) $= \dfrac{N}{L \text{ (mol}^{-1})}$	N = number of particles of X L = Avogadro constant / mol^{-1}

Worked Example 4

Complete the table below.

Substance	Molar mass / g mol^{-1}	Mass / g	Amount of substance / mol	Number of molecules	Amount of atoms / mol	Number of atoms	Amount of ions / mol	Number of ions
Water, H_2O		**36.0**						
Ethane, C_2H_6			**1.50**					
Sodium chloride, $NaCl$								**3.01 x 10^{24}**

4	Empirical and Molecular Formulae

4.1 Definitions

- The empirical formula of a compound is the formula that shows the simplest whole number ratio of the atoms of different elements present in one molecule or formula unit of the compound.
- The molecular formula of a compound is the formula that shows the actual number of atoms of each element present in one molecule of the compound.
- Examples

Compound	Molecular formula	Empirical formula
methane	CH_4	CH_4
ethene	C_2H_4	CH_2
propene	C_3H_6	CH_2
cyclohexane	C_6H_{12}	CH_2

Note:

- It is possible for a compound to have its empirical formula being the same as its molecular formula.
- The molecular formula is always a multiple of the empirical formula.

- The empirical formula of a compound may be calculated from experimental data obtained from combustion analysis or elemental analysis.
- The molecular formula can be determined from the empirical formula, provided the molar mass or the relative molecular mass of the compound is known.

4.2 Calculations using combustion data

Worked Example 5

0.500 g of an organic compound **X** containing carbon, hydrogen and oxygen gave on complete combustion 0.6875 g of CO_2 and 0.5625 g of H_2O.

(a) Determine the empirical formula of **X**.
(b) If the relative molecular mass of **X** is 32.0, determine its molecular formula.

Solution

(a)

Element	C	H	O
Mass in 0.500 g of X / g	$\frac{0.6875}{12.0+(2)(16.0)}\times 12.0$ $=0.1875$	$\frac{0.5625}{(2)(1.0)+16.0}\times 2\times 1.0$ $=0.0625$	0.500 - 0.1875 - 0.0625 =0.250
Amount / mol	$\frac{0.1875}{12.0}=0.01563$	$\frac{0.0625}{1.0}=0.0625$	$\frac{0.250}{16.0}=0.01563$
Molar ratio	1	4	1

Hence the empirical formula of **X** is CH_4O.

(b) Let molecular formula of **X** be $(CH_4O)_n$.

$$M_r \text{ of } (CH_4O)_n = 32.0$$
$$[12.0+4(1.0)+16.0]\,n = 32.0$$
$$n = 1$$

Hence molecular formula of **X** is CH_4O.

4.3 Calculations using percentage composition by mass

Worked Example 6

The formula of a complex salt Q is $NH_4[Cr(SCN)_x(NH_3)_y]$ and analysis produced the following composition by mass: Cr, 15.5%; S, 38.1%; N, 29.2%.

Calculate the values of x and y in the formula for Q.

Solution

Element	Cr		S		N
% composition	15.5		38.1		29.2
Molar mass / g mol^{-1}	52.0		32.1		14.0
Amount in 100 g of Q / mol					
Molar ratio		:		:	
Molar ratio based on the formula	1	:	x	:	1 + x + y

Hence, x = ____ 1 + x + y = ____

y = ____

4.4 Some important points to take note of

- In the calculation of empirical formula from experimental data, it is a common procedure to round off figures to the corresponding nearest whole numbers in order to get the simplest ratio. The rounding off procedure is usually justified, taking into consideration the possibility of experimental errors.
- However, when figures such as those listed below are encountered, the usual rounding off process may lead to an incorrect empirical formula. Great care should be exercised when these figures are obtained. These figures are usually multiplied by a factor in order to get the correct simplest ratio.

  Note: Hypothetical data are used here.

Ratio calculated	**Incorrect** simplest ratio deduced	**Actual** simplest ratio	Remarks
1.51 : 3.07 : 1.00	2 : 3 : 1	3 : 6 : 2	1.5 = 3/2 (multiply by a factor 2)
1.33 : 3.07 : 1.00	1 : 3 : 1	4 : 9 : 3	1.33 = 4/3 (multiply by a factor 3)
1.25 : 3.07 : 1.00	1 : 3 : 1	5 : 12 : 4	1.25 = 5/4 (multiply by a factor 4)
1.20 : 2.00 : 1.00	1 : 2 : 1	6 : 10 : 5	1.2 = 6/5 (multiply by a factor 5)

- Stoichiometry is the relationship between the amounts of reactants and products in a chemical reaction. It has to do with how much one substance will react with another.
- The amounts of substances undergoing reaction, as given by a balanced chemical equation, are called the stoichiometric amounts.
- The law of conservation of matter is a fundamental principle in stoichiometry. Mass is conserved in all chemical reactions. Every atom of every element must be accounted for since they are not destroyed nor created, but rearranged.

E.g. in the decomposition of $CaCO_3$,

$$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$$

mass of original sample of $CaCO_3(s)$ = the mass of $CO_2(g)$ produced + mass of $CaO(s)$ residue

5.1 Quantitative information from a balanced equation

- Consider the complete combustion of propane gas to produce carbon dioxide and water:

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

- This balanced equation can be used to deduce the following information:

(a) amounts of reactants and products
1 mole of C_3H_8 reacts with 5 moles of O_2 to produce 3 moles of CO_2 and 4 moles of H_2O.

(b) masses of reactants and products
44.0 g of C_3H_8 react with 160.0 g of O_2 to produce 132.0 g of CO_2 and 72.0 g of H_2O.

Worked Example 7

Diallyl sulfide, $H_2C{=}CH–CH_2–S–CH_2–CH{=}CH_2$, can be isolated from garlic.

Which statements about diallyl sulfide (M_r = 114.1) on complete combustion are correct?

1 0.10 g of diallyl sulfide reacts with 0.13 g of O_2.
2 0.10 g of diallyl sulfide produces 0.23 g of CO_2.
3 0.10 g of diallyl sulfide produces 0.876 mol of SO_2.

A 1, 2 and 3 **B** 1 and 2 only **C** 2 and 3 only **D** 2 only

Solution

$$H_2C{=}CH–CH_2–S–CH_2–CH{=}CH_2 + \frac{19}{2}O_2 \rightarrow 6CO_2 + SO_2 + 5H_2O$$

Use your EAR

Step 1: Write a balanced EQUATION *(so that you can figure out the mole ratio between the reactants and products)*
Step 2: Calculate AMOUNT
Step 3: Compare mole RATIO for further calculations

$$\text{Amount of diallyl sulfide in 0.10 g} = \frac{0.10}{114.1} = \underline{8.764 \times 10^{-4}\ \text{mol}} = \text{Amount of } SO_2$$

$$\text{Amount of } O_2 \text{ reacted} = \left(\frac{19}{2}\right)\left(8.764 \times 10^{-4}\right) = 8.326 \times 10^{-3}\ \text{mol}$$

$$\text{Mass of } O_2 \text{ reacted} = 8.326 \times 10^{-3} \times (2 \times 16.0) = \underline{0.266\ \text{g}}$$

$$\text{Amount of } CO_2 \text{ produced} = 6\left(8.764 \times 10^{-4}\right) = 5.258 \times 10^{-3}\ \text{mol}$$

$$\text{Mass of } CO_2 \text{ produced} = 5.258 \times 10^{-3} \times (12.0 + 2 \times 16.0) = \underline{0.231\ \text{g}}$$

5.2 Limiting reagent

- In carrying out chemical reactions, the reactants present may not be in stoichiometric amounts.
- One or more reactants may be in excess of that theoretically needed for complete reaction. The deficient reactant is called the limiting reagent and is consumed completely in the reaction.
- The maximum or theoretical yield of a product is thus determined based on the limiting reagent.

Worked Example 8

Consider the following reaction: $Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$.
What will be the amount of H_2 produced when 3.0 moles of Mg(s) are added to 4.0 moles of HC*l*(aq)?

Solution

Amount of HC*l*(aq) used = 4.0 mol
Amount of Mg required to react with 4.0 mol of HC*l* = (½)(Amount of HC*l*) = (½)(4) = 2.0 mol
Since the amount of Mg available > the amount of Mg required, Mg is present in excess.
⇒ The limiting reagent is HC*l*(aq).

Hence, amount of H_2 produced = (½)(Amount of HC*l* reacted) = (½)(4.0 mol) = **2.00 mol**

5.3 Percentage yield

- The mass of a product formed in a chemical reaction is called the yield.
- The theoretical yield of a product refers to the mass of product calculated from the chemical equation based on the amount of limiting reagent used.
- The actual yield refers to the mass of product that is actually obtained in the reaction during an experiment. It is usually less than the theoretical yield.
- The percentage yield relates the actual yield to the theoretical yield and is expressed as follows:

$$\text{percentage yield} = \frac{\text{actual yield (g)}}{\text{theoretical yield (g)}} \times 100\%$$

Worked Example 9

Ethanol can undergo esterification with ethanoic acid in the presence of concentrated sulfuric acid to yield ethyl ethanoate:

$$\underset{\text{ethanol}}{CH_3CH_2OH(l)} + \underset{\text{ethanoic acid}}{CH_3COOH(l)} \rightarrow \underset{\text{ethyl ethanoate}}{CH_3COOCH_2CH_3(l)} + H_2O(l)$$

(a) In an experiment, 23.0 g of ethanol (the limiting reagent) were used for esterification. Calculate the theoretical yield of ethyl ethanoate.

(b) If 30.0 g of ethyl ethanoate were obtained in the above experiment, determine the percentage yield.

Solution

(a) Amount of ethyl ethanoate produced

= Amount of ethanol used

$$= \frac{23.0}{2(12.0) + 6(1.0) + 16.0} = 0.500 \text{ mol}$$

∴ Theoretical yield of ethyl ethanoate
= 0.500 [4(12.0) + 8(1.0) + 2(16.0)] = 44.0 g

(b)

$$\%\text{Yield} = \frac{30.0}{44.0} \times 100\% = 68.2\%$$

6	Reacting Volumes of Gases

6.1 Avogadro's hypothesis

- Avogadro's hypothesis states that:

Equal volumes of all gases under the same conditions of temperature and pressure contain the same number of molecules.

Note: Since the hypothesis has been used successfully for nearly two centuries, it is often called Avogadro's Law.

- It follows from Avogadro's hypothesis that if equal volumes of gases contain equal numbers of molecules, then the volume occupied by 1 mole of a gas (i.e. containing 6.02 x 10^{23} gas particles) must be the same for all gases.

Number of particles	6.02 x 10^{23} CO_2 molecules	6.02 x 10^{23} O_2 molecules	6.02 x 10^{23} Ne atoms
Amount / mol	1.0	1.0	1.0
Mass / g	44.0	32.0	20.2
Volume of gas at same temperature and pressure / dm^3	V	V	V

- This volume occupied by 1 mole of a gas (which contains 6.02 x 10^{23} gas particles) is termed the molar volume of the gas and is the same for all gases under the same conditions of temperature and pressure.

6.2 Molar volume of a gas

- Note the following data which are commonly used in calculations:

The molar volume (symbol: V_m) of a gas is the volume occupied by 1 mole of the gas.	V_m = **22.7 dm^3 mol^{-1} at s.t.p.** where s.t.p. is expressed as 10^5 Pa (1 bar) and 273 K (0°C)
	V_m = **24 dm^3 mol^{-1} at r.t.p.** where r.t.p. is expressed as 101325 Pa (1 atm) and 293 K (20°C)

- The amount of a gas X can be determined from its volume and molar volume measured under the same conditions of temperature and pressure.

$$\text{Amount of gas X (mol)} = \frac{\text{volume of X } (dm^3)}{\text{molar volume of X } (dm^3\,mol^{-1})}$$

Conversion of units for volume

1 cm^3	=	10^{-3} dm^3	=	10^{-6} m^3
1 dm^3	=	1000 cm^3	=	10^{-3} m^3
1 m^3	=	1000 dm^3	=	10^6 cm^3

- Using the previous example: $C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(l)$

	$C_3H_8(g)$	$5O_2(g)$	$3CO_2(g)$
Vol. of gas at s.t.p. / dm^3	22.7	5 x 22.7	3 x 22.7
Vol. of gas at r.t.p. / dm^3	24	5 x 24	3 x 24
Vol. of gas at same T and P / dm^3	10		
	reacted	reacted	produced

Worked Example 10

In the preparation of hydrogen chloride by the reaction

$$NaCl(s) + H_2SO_4(l) \longrightarrow HCl(g) + NaHSO_4(s),$$

what mass of sodium chloride is required for the production of 10.0 dm^3 of hydrogen chloride (at s.t.p.)?

Solution

Amount of HC*l*(g) produced = $\frac{10.0}{22.7}$ = 0.4405 mol

Amount of NaC*l*(s) required = Amount of HC*l*(g) produced = 0.4405 mol

Molar mass of NaC*l* = 23.0 + 35.5 = 58.5 g mol^{-1}

Hence, mass of NaC*l* required = (0.4405)(58.5) = **25.8 g**

6.3 Calculations involving combustion of hydrocarbons

- General equation for the complete combustion of a hydrocarbon:

$C_xH_y(g) + (x+\frac{y}{4})\ O_2(g) \longrightarrow x\ CO_2(g) + \frac{y}{2} H_2O(l)$	**Note:** • The state symbol for C_xH_y can either be (g) or (l) depending on the hydrocarbon analysed. • $H_2O(l)$ if volume measurement is done at room temperature. • $H_2O(g)$ if volume measurement is done at a temperature equal to or greater than 100 °C.
$C_xH_y(g) + (x+\frac{y}{4})\ O_2(g) \longrightarrow x\ CO_2(g) + \frac{y}{2} H_2O(g)$	

- **Note:** For gaseous reactions, the molar ratio shown in the balanced equation also indicates the volume ratio of the reactants and products.

	$C_xH_y(g)$ +	$(x+\frac{y}{4})\ O_2(g)$ ⟶	$x\ CO_2(g)$ +	$\frac{y}{2} H_2O(g)$
Molar ratio	1 :	$(x+\frac{y}{4})$:	x :	$\frac{y}{2}$
Volume ratio	1 :	$(x+\frac{y}{4})$:	x :	$\frac{y}{2}$

Worked Example 11

Skunks are mammals best known for their ability to excrete a strong, foul-smelling odour in order to ward off potential attackers. Methanethiol, CH_3SH, is a foul-smelling chemical which is found in the sprays of skunks. A 20 cm^3 sample of methanethiol was exploded with 80 cm^3 of oxygen to produce only CO_2, SO_2 and H_2O.

What would be the final volume of the resultant mixture of gases when cooled to room temperature and pressure?

Solution

	$CH_3SH(g)$ +	$3O_2(g)$ →	$CO_2(g)$ +	$SO_2(g)$ +	$2H_2O(l)$
Initial volume /cm^3	20	80	0	0	
Change in volume / cm^3	– 20	– 3(20)	+ 20	+ 20	
Final volume / cm^3	0	20	20	20	

Volume of unreacted O_2 = 80 – 60 = 20 cm^3

Final volume of the resultant mixture of gases = 20 + 20 + 20 = **60 cm^3**

Worked Example 12

10 cm^3 of a hydrocarbon was burnt in 100 cm^3 (an excess) of oxygen gas. After reaction, the volume of the gaseous mixture was found to be 80 cm^3 when cooled to room temperature. Upon passing the gaseous mixture through aqueous potassium hydroxide, its volume decreased to 50 cm^3.

What is the molecular formula of the unknown hydrocarbon?

Solution

$$\underset{\textbf{10 cm}^3}{C_xH_y} + \underset{\textbf{100 cm}^3\text{ (excess)}}{O_2(g)} \xrightarrow{\text{burnt \& cooled}} \underset{\textbf{80 cm}^3}{CO_2(g) + \text{unreacted } O_2(g)} + H_2O(l)$$

$$\underset{\textbf{80 cm}^3}{CO_2(g) + \text{unreacted } O_2(g)} \xrightarrow{KOH(aq)} \underset{\textbf{50 cm}^3}{\text{unreacted } O_2(g)}$$

KOH reacts with acidic species, e.g. CO_2, absorbing the acidic gas and hence lowering the total volume of gas.

Let the hydrocarbon be C_xH_y.
After reaction, the gaseous mixture contained CO_2 and unreacted O_2.
When the gaseous mixture was passed through KOH(aq), only the acidic CO_2 gas reacted with KOH.

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

Volume of unreacted O_2 = 50 cm^3
Volume of CO_2 produced = 80 – 50 = 30 cm^3
Volume of O_2 reacted = 100 – 50 = 50 cm^3

	$C_xH_y(g)$	+ $(x+\frac{y}{4})\ O_2(g)$	$\rightarrow x\ CO_2(g)$	+ $\frac{y}{2} H_2O(l)$
Volume of gas / cm^3	10			(negligible)
Volume ratio	:	:		
Molar ratio	:	:		
Molar ratio (from equation)	1 :	$(x+\frac{y}{4})$:	x	

Hence, x = ___ and $(x+\frac{y}{4})$ = ___ $\therefore$ the molecular formula of the hydrocarbon is ________.

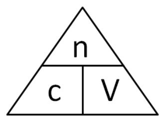
$\Rightarrow$ y = ___

7 Concentration of a Solution

7.1 Definitions

- A solution is a homogeneous mixture of two or more substances. The substance which is in greater quantity is the solvent and the other substance is called the solute.
- The term concentration is used to designate the amount (or mass) of solute dissolved in a given quantity of solvent or solution.
- The concentration of a solute X in a solution can be expressed in various units.

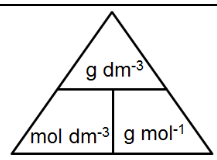
(a) molar concentration (units: mol dm^{-3})

concentration of X (mol dm^{-3}) = $\dfrac{\text{amount of X (mol)}}{\text{volume of solution (dm}^3\text{)}}$	
or $[\,X\,] = \dfrac{n_X}{V}$	[X] = concentration of X (in mol dm^{-3})
or $c_X = \dfrac{n_X}{V}$	c_X = concentration of X (in mol dm^{-3})

(b) mass concentration (units: g dm^{-3})

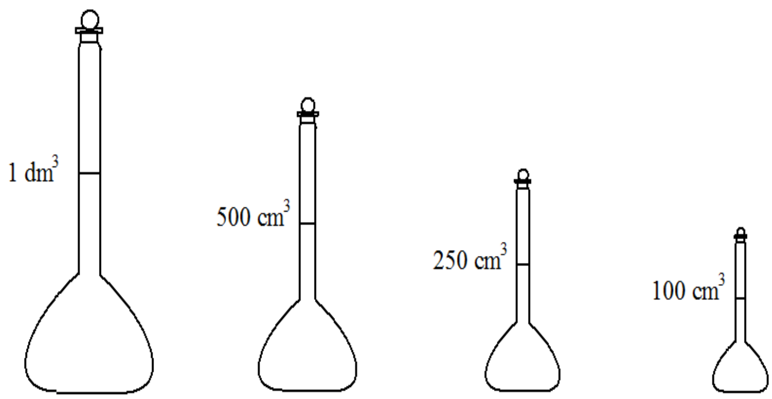
concentration of X (g dm^{-3}) = $\dfrac{\text{mass of X (g)}}{\text{volume of solution (dm}^3\text{)}}$

- The above two concentration terms in different units are related by the following expression:

concentration of X in mol dm^{-3} = $\dfrac{\text{concentration of X in g dm}^{-3}}{\text{molar mass of X in g mol}^{-1}}$	
or mass concentration (g dm^{-3}) = molar concentration (mol dm^{-3}) x molar mass (g mol^{-1})	

7.2 Volumetric/ graduated flasks

- A solution of a particular concentration is usually prepared by dissolving the required amount of solute in a solvent (usually deionised water) in a volumetric/ graduated flask of suitable capacity.
- The volume of solvent added to the flask is such that it makes up the required volume of the solution.



7.3 Calculations involving concentration of a solution

Worked Example 13

Calculate the mass of $Mg(NO_3)_2.6H_2O$ required to prepare a 250 cm^3 solution containing 0.250 mol dm^{-3} of nitrate(V) ions.

Solution

$$Mg(NO_3)_2.6H_2O(s) + aq \rightarrow Mg^{2+}(aq) + 2NO_3^-(aq) + 6H_2O(l)$$

$$\text{Amount of } NO_3^- \text{ ions} = \frac{250}{1000} \times 0.250 = 6.25 \times 10^{-2} \text{ mol}$$

$$\text{Amount of } Mg(NO_3)_2.6H_2O \text{ required} = \frac{1}{2} \times 6.25 \times 10^{-2} = 3.125 \times 10^{-2} \text{ mol}$$

$$\text{Mass of } Mg(NO_3)_2.6H_2O \text{ required} = 3.125 \times 10^{-2} \times [24.3 + 2(14.0 + 3 \times 16.0) + 6(2 \times 1.0 + 16.0)]$$

$$= \underline{8.01 \text{ g}}$$

Worked Example 14

6.45 g of sodium thiosulfate pentahydrate ($Na_2S_2O_3.5H_2O$) is dissolved in deionised water and the volume made up to 250 cm^3 in a volumetric flask.

(a) What is the concentration of the sodium thiosulfate ($Na_2S_2O_3$) solution prepared?
(b) What is the concentration of the thiosulfate ions (i.e. $S_2O_3^{2-}$ ions) in the solution?
(c) Determine the concentration in g dm^{-3} of the thiosulfate ions in the solution.
(d) If 40.00 cm^3 of the solution is removed and made up to 250 cm^3 in another volumetric flask using deionised water, what is the concentration of thiosulfate ions in the new solution?

Solution

(a) $Na_2S_2O_3.5H_2O(s) + aq. \rightarrow Na_2S_2O_3(aq) + 5H_2O(l)$

$$\text{Amount of } Na_2S_2O_3.5H_2O = \frac{6.45}{2 \times 23.0 + 2 \times 32.1 + 3 \times 16.0 + 5(2 \times 1.0 + 16.0)} = 2.599 \times 10^{-2} \text{ mol}$$

$$\text{Amount of } Na_2S_2O_3 = \text{Amount of } Na_2S_2O_3.5H_2O = 2.599 \times 10^{-2} \text{ mol}$$

$$\text{Concentration of } Na_2S_2O_3 = \frac{2.599 \times 10^{-2}}{250 \times 10^{-3}} = \underline{0.104 \text{ mol dm}^{-3}}$$

(b) $Na_2S_2O_3(aq) \rightarrow 2Na^+(aq) + S_2O_3^{2-}(aq)$

$$\text{Concentration of } S_2O_3^{2-} = \text{Concentration of } Na_2S_2O_3 = \underline{0.104 \text{ mol dm}^{-3}}$$

(c) Concentration of $S_2O_3^{2-}$ in g dm^{-3} = 0.104 × (2 x 32.1 + 3 x 16.0) = $\underline{11.7 \text{ g dm}^{-3}}$

(d) Amount of $S_2O_3^{2-}$ ions in 250 cm^3 of diluted solution

= Amount of $S_2O_3^{2-}$ ions in 40.00 cm^3 of original solution

$$= \frac{40.00}{1000} \times 0.104 = 4.16 \times 10^{-3} \text{ mol}$$

Concentration of $S_2O_3^{2-}$ in new solution

$$= \frac{4.16 \times 10^{-3}}{250 \times 10^{-3}} = \underline{0.0166 \text{ mol dm}^{-3}}$$

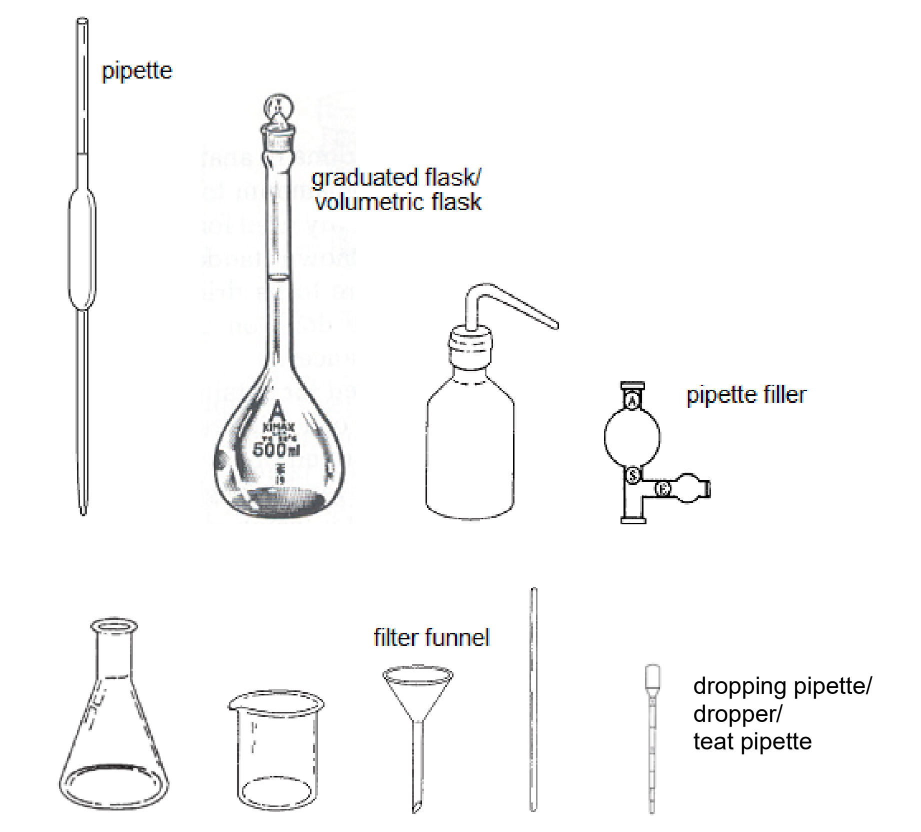
Alternative presentation of working for (d):

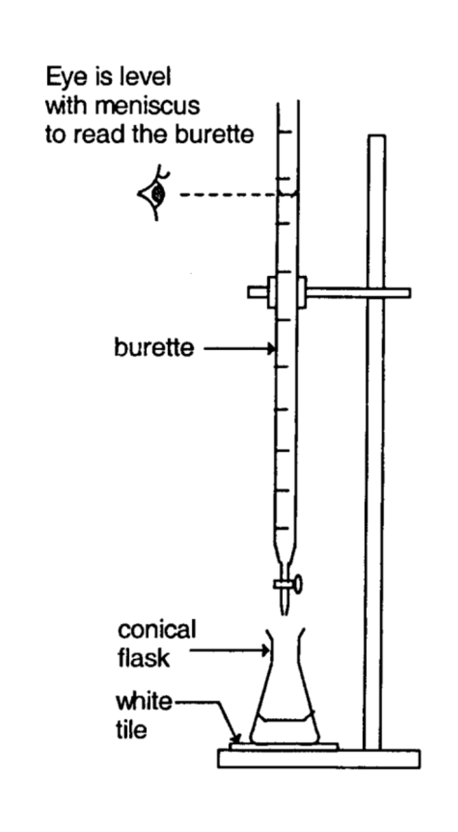
Amount of $S_2O_3^{2-}$ ions in 250 cm^3 of diluted solution	=	Amount of $S_2O_3^{2-}$ ions in 40.00 cm^3 of original solution
c_1V_1	=	c_2V_2
$c_1\left(\frac{250}{1000}\right)$	=	$0.104\left(\frac{40.00}{1000}\right)$
c_1	=	0.0166 mol dm^{-3}
[$S_2O_3^{2-}$] in the diluted solution	=	$\underline{0.0166 \text{ mol dm}^{-3}}$

8 Acid-Base Titrations

8.1 Volumetric Analysis

- Volumetric analysis (or titrimetric analysis) is a method of quantitative analysis which depends essentially on the accurate measurements of the volumes of two solutions which react together completely.
- In volumetric analysis, a standard solution (i.e. a solution of known concentration) is used to determine the concentration of another solution.
- This is done through a titration process which involves the gradual addition of one solution (from a burette) to a fixed volume of another solution (in a conical flask) until stoichiometric amounts of the two reactants have reacted.
  - The solution that is placed in the burette is sometimes referred to as the **titrant**.
  - The solution that is pipetted into the conical flask is less commonly referred to as the **titrand**.
- In practice, the completion of a titration is usually detected by a distinct colour change brought about by the use of a suitable indicator. The point at which this distinct colour change occurs is called the end-point of the titration.
- In an acid-base titration, the indicator used is usually added in a small quantity. Common indicators include methyl orange, screened methyl orange, thymol blue and thymolphthalein. Each indicator has a pH range over which it changes colour.
- The common apparatus used in a titration are shown below.





8.2 Acid-base titrations

- Acid-base titrations are carried out in order to establish the stoichiometric amounts of acid and base which are required to neutralise each other.

- Basicity of an acid

	Number of H atom ionisable (as H^+ ion) per molecule	Examples
monobasic (or monoprotic) acid	1	HCl, HNO_3 and CH_3COOH
dibasic acid (or diprotic) acid	2	H_2SO_4, HOOC–COOH
tribasic acid (or triprotic) acid	3	H_3PO_4

8.3 Calculations involving acid-base titration

Worked Example 15

In a titration experiment, 20.0 cm^3 of 0.200 mol dm^{-3} NaOH reacted with 32.00 cm^3 of H_2SO_4 solution. Calculate the concentration of H_2SO_4 in (a) mol dm^{-3}, and (b) g dm^{-3}.

Solution

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

$$\text{Amount of NaOH reacted} = \frac{20.0}{1000} \times 0.200 = 4.00 \times 10^{-3} \text{ mol}$$

$$\text{Amount of } H_2SO_4 \text{ reacted} = \left(\frac{1}{2}\right)(\text{Amount of NaOH reacted}) = \frac{1}{2} \times 4.00 \times 10^{-3} = 2.00 \times 10^{-3} \text{ mol}$$

$$\text{Concentration of } H_2SO_4 = \frac{2.00 \times 10^{-3}}{32.00 \times 10^{-3}} = \underline{0.0625 \text{ mol dm}^{-3}}$$

(b) Concentration of H_2SO_4 = (0.0625)(2 X 1.0 + 32.1 + 4 X 16.0) = $\underline{6.13 \text{ g dm}^{-3}}$

Worked Example 16

An acid solution contains 25.2 g of $(COOH)_2.xH_2O$ per dm^3. 50.0 cm^3 of this solution needed 40.00 cm^3 of 0.500 mol dm^{-3} KOH(aq) for complete neutralisation. Calculate the value of x.

Solution:

$$\text{Concentration of } (COOH)_2.xH_2O \text{ solution} = \frac{25.2}{(90.0 + 18.0x)} \text{ mol dm}^{-3}$$

$(COOH)_2.xH_2O(aq) + 2KOH(aq) \rightarrow (COO^-K^+)_2(aq) + (x+2)H_2O(l)$

$$\text{Amount of KOH reacted} = \left(\frac{40.0}{1000}\right)(0.500) = 0.0200 \text{ mol}$$

Amount of $(COOH)_2.xH_2O$ reacted
= (½)(Amount of KOH reacted)
= (½)(0.0200)
= 0.0100 mol

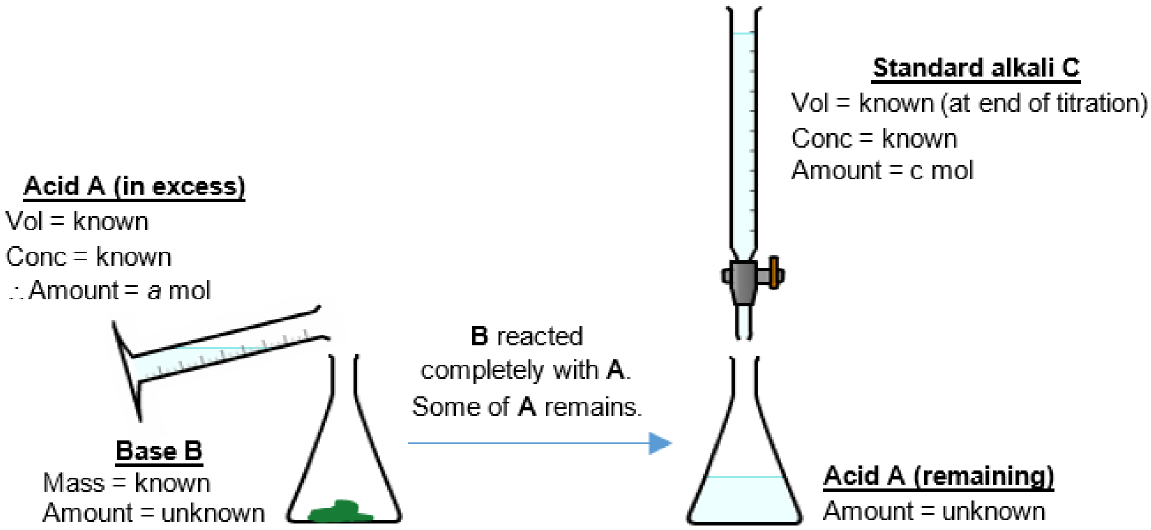
$$[(COOH)_2.xH_2O] = \frac{0.0100}{50.0 \times 10^{-3}} = 0.200 \text{ mol dm}^{-3}$$

Hence, $[(COOH)_2.xH_2O] =$

$x =$

8.4 Back titration

- Back titrations are usually employed when the determination of the amount of a substance poses some difficulty in the direct titration method, e.g. solid substances ($CaCO_3$) where the end-point is difficult to detect, and volatile substances (ammonia, iodine) where inaccuracy arises due to loss of substance during titration.
- In a typical acid-base back titration, the following steps are involved:
    - A known excess of an acid A is added to a quantity of a base B (or vice versa).
    - Upon mixing, the base B reacts completely while only some of the acid A reacts.
    - The acid A remaining is then titrated with a standard alkali C and its amount determined.
    - From the results, the amount of acid A which has reacted with the base B can be found and the amount of base can then be calculated.



Worked Example 17

A 3.00 g mixture of calcium carbonate and calcium chloride was added to 50.0 cm^3 of 1.00 mol dm^{-3} hydrochloric acid. The resulting solution was made up to 250 cm^3 in a graduated flask with deionised water. When 25.0 cm^3 of this solution was titrated with 0.100 mol dm^{-3} sodium hydroxide solution, 21.05 cm^3 of sodium hydroxide was required for complete reaction.

Calculate the percentage by mass of calcium carbonate in the given mixture.

Solution

$CaCO_3(s)$ and $CaCl_2(s)$	+	$HCl(aq)$ 50.0 cm^3 1.00 mol dm^{-3}	reaction $\longrightarrow$	some $HCl(aq)$ remained	H_2O $\longrightarrow$	$HCl(aq)$ 250 cm^3 solution	pipette $\longrightarrow$	$HCl(aq)$ 25.0 cm^3	titration $\longrightarrow$	$NaOH(aq)$ 21.05 cm^3 0.100 mol dm^{-3}		

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

$HCl + NaOH \rightarrow NaCl + H_2O$

Amount of NaOH reacted = $(\frac{21.05}{1000})(0.100) = 2.105 \times 10^{-3}$ mol

$HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$

Amount of HCl that reacted with NaOH = 2.105×10^{-3} mol

Amount of HCl in 250 cm^3 of solution = $(\frac{250}{25.0})(2.105 \times 10^{-3}) = 2.105 \times 10^{-2}$ mol

Initial amount of HCl used = $(\frac{50.0}{1000})(1.00) = 5.00 \times 10^{-2}$ mol

Amount of HCl that reacted with $CaCO_3$ = $5.00 \times 10^{-2} - 2.105 \times 10^{-2} = 2.895 \times 10^{-2}$ mol

$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$

Amount of $CaCO_3$ reacted = $(\frac{1}{2})(2.895 \times 10^{-2}) = 1.448 \times 10^{-2}$ mol

Molar mass of $CaCO_3$ = 40.1 + 12.0 + 3 x 16.0 = 100.1 g mol^{-1}

Mass of $CaCO_3$ = $(1.448 \times 10^{-2})(100.1) = 1.449$ g

Hence percentage by mass of $CaCO_3$ in the mixture = $(\frac{1.449}{3.00})(100\%)$ = **48.3%**

Note:

When presenting the working for questions involving calculations,

- show essential steps
- write proper statements / use proper symbols
- use A_r values from the Periodic Table given in the *Data Booklet*
- present the answers for intermediate steps correct to 3 or 4 significant figures
- present the final answer to 3 significant figures unless instructed otherwise
- include relevant units

Accepted statements:	**Not** accepted:
• Amount of **X** = 4.00 mol • Number of moles of **X** = 4.00 mol • n_X = 4.00 mol	• Amount of **X** = 4.00 **mols**

Group																	
1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
							1 H hydrogen 1.0										2 He helium 4.0
3 Li lithium 6.9	4 Be beryllium 9.0		**Key** atomic number atomic symbol name relative atomic mass									5 B boron 10.8	6 C carbon 12.0	7 N nitrogen 14.0	8 O oxygen 16.0	9 F fluorine 19.0	10 Ne neon 20.2
11 Na sodium 23.0	12 Mg magnesium 24.3											13 Al aluminium 27.0	14 Si silicon 28.1	15 P phosphorus 31.0	16 S sulfur 32.1	17 Cl chlorine 35.5	18 Ar argon 39.9
19 K potassium 39.1	20 Ca calcium 40.1	21 Sc scandium 45.0	22 Ti titanium 47.9	23 V vanadium 50.9	24 Cr chromium 52.0	25 Mn manganese 54.9	26 Fe iron 55.8	27 Co cobalt 58.9	28 Ni nickel 58.7	29 Cu copper 63.5	30 Zn zinc 65.4	31 Ga gallium 69.7	32 Ge germanium 72.6	33 As arsenic 74.9	34 Se selenium 79.0	35 Br bromine 79.9	36 Kr krypton 83.8
37 Rb rubidium 85.5	38 Sr strontium 87.6	39 Y yttrium 88.9	40 Zr zirconium 91.2	41 Nb niobium 92.9	42 Mo molybdenum 95.9	43 Tc technetium –	44 Ru ruthenium 101.1	45 Rh rhodium 102.9	46 Pd palladium 106.4	47 Ag silver 107.9	48 Cd cadmium 112.4	49 In indium 114.8	50 Sn tin 118.7	51 Sb antimony 121.8	52 Te tellurium 127.6	53 I iodine 126.9	54 Xe xenon 131.3
55 Cs caesium 132.9	56 Ba barium 137.3	57–71 lanthanoids	72 Hf hafnium 178.5	73 Ta tantalum 180.9	74 W tungsten 183.8	75 Re rhenium 186.2	76 Os osmium 190.2	77 Ir iridium 192.2	78 Pt platinum 195.1	79 Au gold 197.0	80 Hg mercury 200.6	81 Tl thallium 204.4	82 Pb lead 207.2	83 Bi bismuth 209.0	84 Po polonium –	85 At astatine –	86 Rn radon –
87 Fr francium –	88 Ra radium –	89–103 actinoids	104 Rf rutherfordium –	105 Db dubnium –	106 Sg seaborgium –	107 Bh bohrium –	108 Hs hassium –	109 Mt meitnerium –	110 Ds darmstadtium –	111 Rg roentgenium –	112 Cn copernicium –	113 Nh nihonium –	114 Fl flerovium –	115 Mc moscovium –	116 Lv livermorium –	117 Ts tennessine –	118 Og oganesson –

lanthanoids	57 La lanthanum 138.9	58 Ce cerium 140.1	59 Pr praseodymium 140.9	60 Nd neodymium 144.2	61 Pm promethium –	62 Sm samarium 150.4	63 Eu europium 152.0	64 Gd gadolinium 157.3	65 Tb terbium 158.9	66 Dy dysprosium 162.5	67 Ho holmium 164.9	68 Er erbium 167.3	69 Tm thulium 168.9	70 Yb ytterbium 173.1	71 Lu lutetium 175.0
actinoids	89 Ac actinium –	90 Th thorium 232.0	91 Pa protactinium 231.0	92 U uranium 238.0	93 Np neptunium –	94 Pu plutonium –	95 Am americium –	96 Cm curium –	97 Bk berkelium –	98 Cf californium –	99 Es einsteinium –	100 Fm fermium –	101 Md mendelevium –	102 No nobelium –	103 Lr lawrencium –

molar volume of gas	V_m = 22.7 dm^3 mol^{-1} at s.t.p. V_m = 24 dm^3 mol^{-1} at r.t.p. (where s.t.p. is expressed as 10^5 Pa [1 bar] and 273 K [0 °C], r.t.p. is expressed as 101325 Pa [1 atm] and 293 K [20 °C])