Raffles Institution
Year 5 H2 Chemistry 2022
Lecture Notes 1a – 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

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 http:///www.chemguide.co.uk

1.1 The sub-atomic particles

- An <u>atom</u> is the smallest part of an element which can ever exist, whereas a <u>molecule</u> 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 — <u>protons</u>, <u>neutrons</u> and <u>electrons</u>.

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 <u>proton number (or atomic number)</u> of an element is the <u>number of protons</u> in the nucleus of an atom of that element. • The atomic number determines the identity of an element. 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 <u>nucleon number (or mass number)</u> of an element is the <u>total number of protons and neutrons</u> in the nucleus of an atom of that element. • **Note:** Protons and neutrons are collectively known as <u>nucleons</u> because they are both found in the nucleus.
(c)	nuclide	$^{A}_{Z}X$	• A <u>nuclide</u> 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 nucleon number (or mass number) $\rightarrow$ A proton number (or atomic number) $\rightarrow$ Z $\boxed{^{A}_{Z}X}$ $\leftarrow$ symbol of the element • **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^-

$^{1}_{1}H$ $^{16}_{8}O$ $^{2}_{1}D$

(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	10
No. of protons	1	9	11	9
No. of neutrons	1	8	8	9

(b)

	P	Q^{2-}	R^+	S^-	T^+
No. of electrons	11	11	10	12	13
No. of protons	11	9	11	11	14
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:

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 $^{21}Ne = 20.994 \approx 21.0$
  Relative isotopic mass of $^{35}Cl = 34.97 \approx 35.0$
- 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:

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

$$A_r \text{ of } Cl = \frac{(75.53)(34.97) + (24.47)(36.95)}{75.53 + 24.47} = \underline{\mathbf{35.5}} \text{ (3 s.f.)}$$

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.

$$A_r \text{ of Br} = \frac{(x)(78.918) + (100-x)(80.916)}{x + (100-x)} = 79.904$$

Percentage abundance of ^{79}Br = x = 50.7% (3sf)

Percentage abundance of ^{81}Br = 100 – x = 49.3% (3sf)

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

3.1 The mole and the Avogadro constant

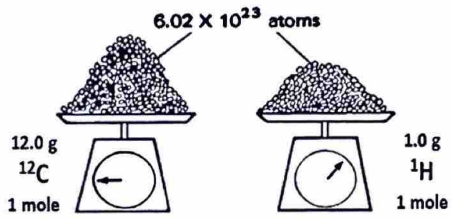
- The mole is the SI (Système International) unit for measuring "amount of substance".

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 6.02×10^{23} carbon atoms.
  $\Rightarrow$ 1 mole of carbon-12 contains 6.02×10^{23} carbon-12 atoms.
- The definition of the mole reads as follows:

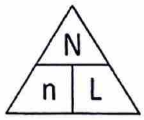
A mole of substance is the amount of that substance which contains 6.02×10^{23} elementary entities of that substance.

Note: The elementary entities must be specified and may be atoms, molecules, ions, electrons, other particles, or other formula units. The quantity of 6.02×10^{23} mol^{-1} is termed the Avogadro constant (Symbol: *L*)



- 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 **6.02×10^{23}** H_2O molecules.
- contains **1 mole** of O atoms and hence contains **6.02×10^{23}** O atoms.
- contains **2 moles** of H atoms and hence contains **$(6.02 \times 10^{23})(2)$** H atoms.
- contains **3 moles** of atoms and hence contains **$(6.02 \times 10^{23})(3)$** atoms.

1 mole of $MgCl_2$
- contains **6.02×10^{23}** formula units of $MgCl_2$.
- contains **1 mole** of Mg^{2+} ions and hence contains **6.02×10^{23}** Mg^{2+} ions.
- contains **2 moles** of Cl^- ions and hence contains **$(6.02 \times 10^{23})(2)$** Cl^- ions.
- contains **3 moles** of ions and hence contains **$(6.02 \times 10^{23})(3)$** ions.

3.2 Molar mass

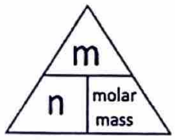
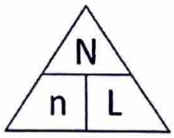
- The molar mass of a substance is the mass of one mole of the substance. It has units of g mol^{-1}.
- Examples:

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 Relationship between amount, mass, molar mass and number of particles

- Take note of the following:

Amount of substance X (mol) = $\dfrac{\text{mass of X (g)}}{\text{molar mass of X (g mol}^{-1}\text{)}}$ or n_X (mol) = $\dfrac{m_X \text{ (g)}}{\text{molar mass of X (g mol}^{-1}\text{)}}$	 m_X = mass of substance X / g n_X = amount of X / mol
Amount of substance X (mol) = $\dfrac{\text{number of particles of X}}{\text{Avogadro constant (mol}^{-1}\text{)}}$ or n_X (mol) = $\dfrac{N}{L \text{ (mol}^{-1}\text{)}}$	 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	18.0	36.0	2.00	1.20×10^{24}	6.00	3.61×10^{24}		
Ethane, C_2H_6	30.0	45.0	1.50	9.03×10^{23}	12.0	7.22×10^{24}		
Sodium chloride, NaCl	23.0 + 35.5 = 58.5	58.5 × 2.50 = 146 (3sf)	$\frac{1}{2}$ × 5.00 = 2.50				5.00	3.01×10^{24}

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	$\frac{15.5}{52.0}$ X 0.298		$\frac{38.1}{32.1}$ = 1.19		$\frac{29.2}{14.0}$ = 2.09
Molar ratio	1	:	4	:	7
Molar ratio based on the formula	1	:	x	:	1 + x + y

Hence, x = 4 1 + x + y = 7

y = 3

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

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

Amount of diallyl sulfide in 0.10 g $= \frac{0.10}{114.1}$

$= \underline{8.764 \times 10^{-4} \text{ mol}}$

$=$ Amount of SO_2

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

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

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

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

Mass of CO_2 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 HCl(aq)?

Solution

Amount of HCl(aq) used = 4.0 mol
Amount of Mg required to react with 4.0 mol of HCl = (½)(Amount of HCl) = (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 HCl(aq).

Hence, amount of H_2 produced = (½)(Amount of HCl 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 6.02×10^{23} gas particles) must be the same for all gases.

Number of particles	6.02×10^{23} CO_2 molecules	6.02×10^{23} O_2 molecules	6.02×10^{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 6.02×10^{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: $C_3H_8(g)$ + $5O_2(g)$ → $3CO_2(g)$ + $4H_2O(l)$

	$C_3H_8(g)$	$5O_2(g)$	$3CO_2(g)$
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	50	30

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 $HCl(g)$ produced $= \frac{10.0}{22.7} = 0.4405$ mol

Amount of $NaCl(s)$ required = Amount of $HCl(g)$ produced = 0.4405 mol

Molar mass of $NaCl$ = 23.0 + 35.5 = 58.5 g mol^{-1}

Hence, mass of $NaCl$ 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.

	$C_xH_y(g)$	+ $(x+\frac{y}{4})\,O_2(g)$	$\longrightarrow x\,CO_2(g)$	+ $\frac{y}{2}H_2O(g)$
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

	$CH_3SH(g)$	+ $3O_2(g)$	$\rightarrow CO_2(g)$	+ $SO_2(g)$	+ $2H_2O(l)$
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?

<u>Solution</u>

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

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

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

	$C_xH_y(g)$ +	$(x+\frac{y}{4})\ O_2(g)$ →	$x\ CO_2(g)$ +	$\frac{y}{2}H_2O(l)$
Volume of gas / cm^3	10 :	50 :	30	(negligible)
Volume ratio	1 :	5 :	3	
Molar ratio	1 :	5 :	3	
Molar ratio (from equation)	1 :	$(x+\frac{y}{4})$:	x	

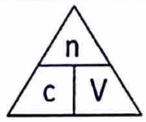
Hence, x = 3 and $(x+\frac{y}{4})$ = 5
$\Rightarrow$ y = 8

∴ the molecular formula of the hydrocarbon is C_3H_8.

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.

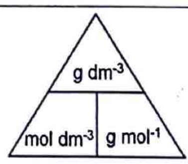
(a) molar concentration (units: mol dm^{-3})

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})

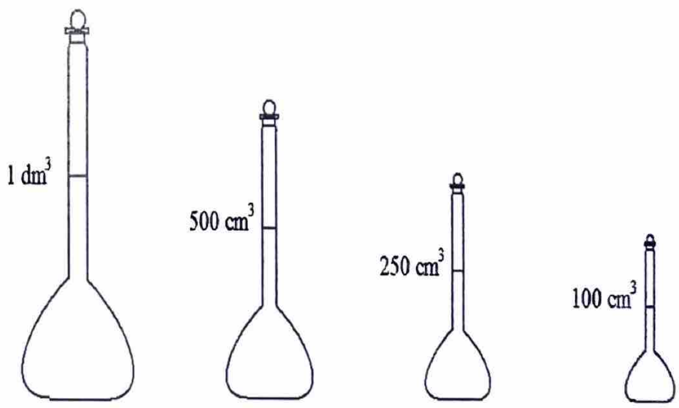
concentration of X (g dm^{-3}) = $\dfrac{\text{mass of X (g)}}{\text{volume of solution (dm}^3\text{)}}$

- 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 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 graduated flask (or volumetric 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 cm³ solution containing 0.250 mol dm⁻³ of nitrate(V) ions.

Solution

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

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

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

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

$= \underline{8.01\ g}$

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 cm³ 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 cm³ of the solution is removed and made up to 250 cm³ in another volumetric flask using deionised water, what is the concentration of thiosulfate ions in the new solution?

Solution

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

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

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

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

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

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

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

(d) Amount of $S_2O_3^{2-}$ ions in 250 cm³ of diluted solution

= Amount of $S_2O_3^{2-}$ ions in 40.00 cm³ of original solution

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

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

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

Alternative presentation of working for (d):

Amount of $S_2O_3^{2-}$ ions in 250 cm³ of diluted solution = Amount of $S_2O_3^{2-}$ ions in 40.00 cm³ of original solution

$$c_1V_1 = c_2V_2$$

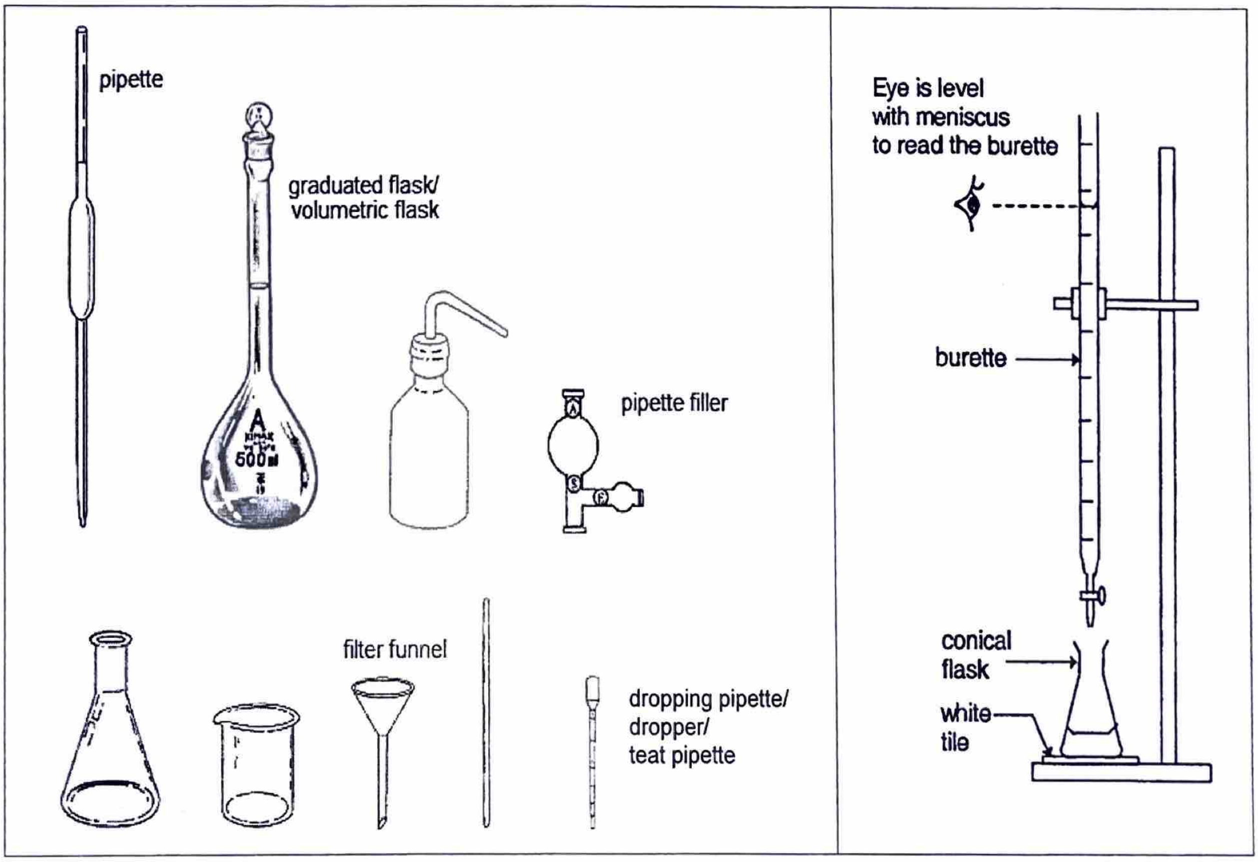
$$c_1\left(\frac{250}{1000}\right) = 0.104\left(\frac{40.00}{1000}\right)$$

$$c_1 = 0.0166\ mol\ dm^{-3}$$

$[S_2O_3^{2-}]$ in the diluted solution $= \underline{0.0166\ mol\ dm^{-3}}$

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 cm³ of 0.200 mol dm⁻³ NaOH reacted with 32.00 cm³ 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 dm³. 50.0 cm³ of this solution needed 40.00 cm³ 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, $[(COOH)_2.xH_2O] = \frac{25.2}{(90.0 + 18.0x)}$

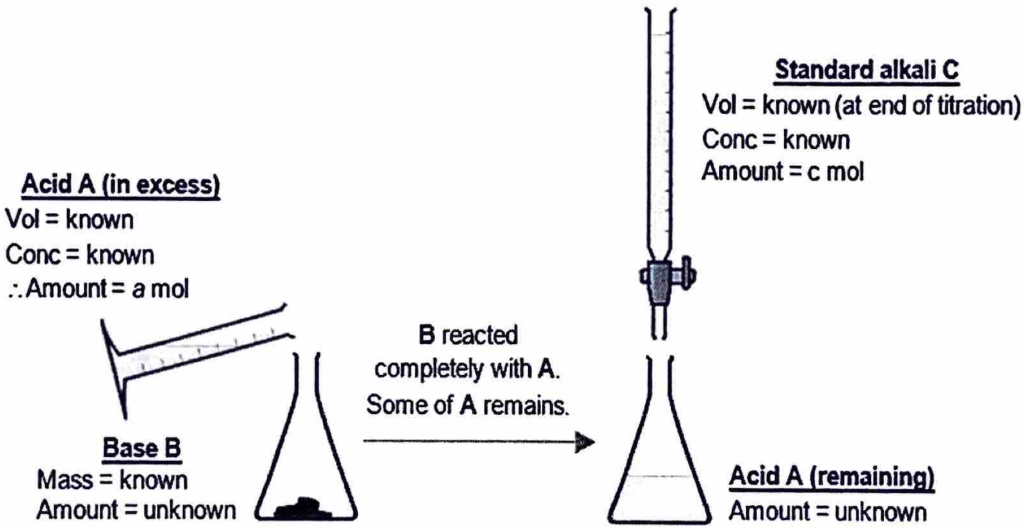
$0.200 = \frac{25.2}{(90.0+18.0x)}$

$90.0 + 18.0x = 126$

x = 2

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.

<u>Solution</u>

$CaCO_3(s)$ 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 $CaCO_3$ = $40.1 + 12.0 + 3 \times 16.0 = 100.1$ 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 = $(\frac{1.449}{3.00})(100\%)$ = **<u>48.3%</u>**

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:	**<u>Not</u>** 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 **mol<u>s</u>**

Raffles Institution
Year 5 H2 Chemistry 2022

Tutorial 1a – The Mole Concept and Stoichiometry

Self-Check Questions

1 2014 RI CT Qn A4 (Modified)
Use of the Data Booklet is relevant to this question.

Which species has both the same number of neutrons and the same number of electrons as an atom of iron–57?

A $^{57}Co^{2+}$ **B** $^{63}Cu^{3+}$ **C** $^{56}Mn^{+}$ **D** $^{59}Ni^{2+}$

2 Naturally occurring silicon is a mixture of three isotopes, ^{28}Si, ^{29}Si and ^{30}Si. The relative atomic mass of silicon is 28.109 and the total percentage abundance of two of the isotopes is 7.8%.

What could be the relative abundance of each of the three isotopes?
[Hint: Which isotope is likely to have % abundance of (100-7.8)%?]

3 Calculate the mass of each of the following:

(a) 3.01×10^{26} oxygen atoms
(b) 6.02×10^{30} bromine molecules
(c) 6.0 mol of CO_3^{2-} ions
(d) C in 200 g of $NaHCO_3$

4 Calculate the number of atoms in each of the following:

(a) 1.00 kg of H_2O **(b)** 1.50 mol of S_8 **(c)** 2.00 dm^3 of CH_4 at s.t.p.

(Note: s.t.p. is expressed as 10^5 Pa [1 bar] and 273 K [0°C])

5 Calculate the amount (in mol) of:

(a) O atoms in 180 g of $CuSO_4.5H_2O$
(b) ions in 4.00 kg of Na_2SO_4
(c) N_2 in 5.00 cm^3 of N_2 at s.t.p.
(d) Fe in 5.00×10^{21} atoms of Fe

6 2014 RI Promo Qn A1
Titanium is a very strong and corrosion-resistant metal with a high melting point. It is extracted from its ores by the following steps:

I. $TiO_2 + C + 2Cl_2 \rightarrow TiCl_4 + CO_2$
II. $TiCl_4 + 2Mg \rightarrow 2MgCl_2 + Ti$

What is the mass of magnesium (in tonnes) needed to extract titanium from 160 tonnes of ore containing 10% by mass of titanium(IV) oxide?
[1 tonne = 10^6 g; M_r of TiO_2 = 79.9]

A 4.87 **B** 9.73 **C** 48.7 **D** 97.3

7 A common ingredient in perfumes is vanillin which occurs widely in nature. Vanillin has the chemical formula, $C_8H_8O_3$, and its scent can be detected with just 2×10^{-11} g in 1 dm^3 of air.

(a) How many moles of vanillin are present in 2×10^{-11} g?

(b) What is the volume of 2×10^{-11} g vanillin vapour at room temperature and pressure?

(c) Your answer to **(b)** is the volume of vanillin vapour in 1 dm^3 of air. How many ppm (parts per million) by volume of vanillin can be detected?
[Note: 1 ppm = 1 part per million = 1 dm^3 in 10^6 dm^3 of air]

8 Sodium peroxide, Na_2O_2, is used in submarines for absorbing atmospheric carbon dioxide and regenerating oxygen. The reaction produces sodium carbonate as a by-product.

(a) Write a balanced equation for this reaction.

(b) Calculate the mass of sodium peroxide needed per day to absorb the CO_2 produced by a crew of eight submariners, each of whom exhales 600 dm^3 of CO_2 (measured at r.t.p.) per day.

9 2017 RI Promo Qn C2(c)(i)
Buckminsterfullerene, C_{60}, is a molecule with 60 carbon atoms arranged in pentagons and hexagons that are similar to those in a football. C_{60} reacts with hydrogen to form hydrofullerenes with the molecular formula $C_{60}H_n$.

When 60 g of C_{60} is reacted with hydrogen gas at 273 K and 1 bar, the volume of hydrogen gas is decreased by 34 dm^3. Find the value of n in the formula of hydrofullerene formed, $C_{60}H_n$.

10 (a) What is the mass of NaOH(s) required to prepare 250 cm^3 of 0.120 mol dm^{-3} NaOH(aq)?

(b) What volume of water must be added to 900 cm^3 of 0.120 mol dm^{-3} NaC*l* solution to dilute it to 0.100 mol dm^{-3}?

(c) What volumes of 0.05 mol dm^{-3} HC*l* and 0.01 mol dm^{-3} HC*l* must be mixed to give 2.0 dm^3 of 0.02 mol dm^{-3} HC*l*?

(d) What is the concentration of excess sulfuric acid in a solution formed by mixing 50 cm^3 of 0.200 mol dm^{-3} sodium hydroxide and 40 cm^3 of 0.250 mol dm^{-3} sulfuric acid?

Suggested solutions to the self-check questions can be found on Ivy. Please use these solutions to check through your working and your final answers. Consult your tutor if you have any questions.

Practice Questions

11 Ascorbic acid (vitamin C), an essential nutrient for humans, is an organic compound containing carbon, hydrogen and oxygen only. When 1.000 g of ascorbic acid was completely burnt in oxygen, 1.500 g of carbon dioxide and 0.405 g of water were formed.

Calculate the empirical formula of ascorbic acid and determine its molecular formula if its relative molecular mass is 176.

12 Many new cars have air bags which rapidly inflate during an accident to protect the front passengers. The air bag contains sodium azide (NaN_3), silicon dioxide and potassium nitrate.

On impact, sodium azide decomposes to sodium and nitrogen. The nitrogen formed inflates the air-bag while the sodium formed reacts with potassium nitrate to form sodium oxide, potassium oxide and additional nitrogen gas which may be used to fill the air bag.

(a) Write an equation for the **(i)** decomposition of sodium azide, and
(ii) reaction between potassium nitrate and sodium.

(b) Calculate the mass of sodium azide needed to inflate an air bag of capacity 60 dm^3 at room temperature and pressure.

13 It is recommended that drinking water contain fluoride (F^-) for prevention of tooth decay. During the purification process, fluoride ions are added to 1000 cm^3 of reservoir water. A 10.00 cm^3 sample of the fluorinated water is taken and transferred to a volumetric flask. Deionised water is added to make a total volume of 100.0 cm^3 of the diluted solution. This diluted solution contains 0.16 ppm of fluoride ions.

(1 ppm of fluoride ions = 1 mg of fluoride ions in 1000 cm^3 of solvent.)

(a) Calculate the mass of F^- ions present in 1000 cm^3 fluorinated reservoir water.

(b) Fluoride is provided by hydrogen hexafluorosilicate, H_2SiF_6. Calculate the mass of H_2SiF_6 added to the reservoir water.

[M_r of H_2SiF_6 = 144.1]

14 A carbonate $\mathbf{M}_2CO_3$ of mass 5.30 g was dissolved in deionised water to produce a 500 cm^3 solution. 25.0 cm^3 of this solution required 26.0 cm^3 of 0.192 mol dm^{-3} HCl for complete neutralisation.

Calculate the concentration of $\mathbf{M}_2CO_3$ in g dm^{-3} and determine the relative atomic mass of **M**.

15 **FA 1** is a solution containing 5.00 g dm^{-3} of a dibasic acid. **FA 2** is a solution containing 5.00 g dm^{-3} of sodium hydroxide. During a titration, 25.0 cm^3 of the dibasic acid solution, **FA 1**, needed 17.00 cm^3 of sodium hydroxide solution, **FA 2**, for complete neutralisation.

Determine the molar mass of the acid.

16 <u>2014 CT Qn B2(b)</u>

Propane, C_3H_8, is a gaseous hydrocarbon fuel that undergoes combustion to produce carbon dioxide and water. 10 cm³ of C_3H_8 and **z** cm³ of N_2 were burned with excess O_2 in a flask and the flask was then cooled to room temperature. $NO_2(g)$ was the only N containing product formed. The gaseous mixture was then passed through 50 cm³ of 0.50 mol dm^{-3} NaOH(aq), which reacted with the acidic gases. The excess NaOH required 24.40 cm³ of 0.35 mol dm^{-3} $H_2SO_4(aq)$ for complete reaction.

(a) Write balanced equations with state symbols for the two combustion reactions, where C_3H_8 and N_2 each reacted with O_2.

(b) Calculate the volume of gaseous product(s) formed from the combustion of C_3H_8.

(c) Calculate the amount of NaOH that has reacted with the gaseous mixture.

(d) NO_2 reacts with NaOH according to the following reaction:

$$2NO_2(g) + 2NaOH(aq) \rightarrow NaNO_2(aq) + NaNO_3(aq) + H_2O(l)$$

Determine the value of **z**.

17 The percentage of nitrogen in foodstuffs gives an indication of how much protein is present. It can be estimated by the following method.

I. The foodstuff is weighed and boiled under reflux with concentrated sulfuric acid for some time. This converts all the nitrogen into ammonium sulfate.

II. An excess of aqueous sodium hydroxide is added, and the mixture again boiled. The liberated ammonia gas is passed into a known excess of dilute hydrochloric acid.

III. The unreacted hydrochloric acid is then titrated with aqueous sodium hydroxide of known concentration.

(a) Write equations for the two reactions occurring in step II.

(b) When 1.00 g of a foodstuff was subjected to the above procedure, and the gas in step II was passed into 50.0 cm³ of 0.100 mol dm^{-3} HC*l*, it was found that only 20.0 cm³ of 0.100 mol dm^{-3} NaOH were needed to neutralise the unreacted acid.

Calculate the percentage by mass of nitrogen in the foodstuff.