06 The Mole Concept and Stoichiometry

1 Relative Masses

Masses are compared with the **carbon-12** standard: one unit is $\frac{1}{12}$ of the mass of one mole of ^{12}C atoms. Relative masses have **no units**.

Term	DEFINITION
Relative isotopic mass	the mass of one mole of an _____ relative to $\frac{1}{12}$ the mass of one mole of ^{12}C atoms
Relative atomic mass, A_r	the _____ mass of one mole of atoms of an element relative to $\frac{1}{12}$ the mass of one mole of ^{12}C atoms
Relative molecular mass, M_r	the average mass of one mole of _____ relative to $\frac{1}{12}$ the mass of one mole of ^{12}C atoms
Relative formula mass, M_r	the average mass of one mole of _____ of a compound relative to $\frac{1}{12}$ the mass of one mole of ^{12}C atoms (ionic compounds)

$$A_r = \frac{\Sigma(\%\ \text{abundance} \times \text{relative isotopic mass})}{100}$$

A_r is rarely a whole number: it is a weighted average over the isotopes.

WORKED EXAMPLE Chlorine is 75.77% ^{35}Cl and 24.23% ^{37}Cl. Find $A_r(Cl)$.

$$A_r = \frac{75.77(35) + 24.23(37)}{100} = ____$$

WORKED EXAMPLE A copper wire has $A_r(Cu) = 63.54$. Find the percentage abundances of ^{63}Cu and ^{65}Cu.

Let ^{63}Cu be x%: $63.54 = \frac{63x + 65(100 - x)}{100}$ → x = _____, so ^{63}Cu = 73.0% and ^{65}Cu = _____

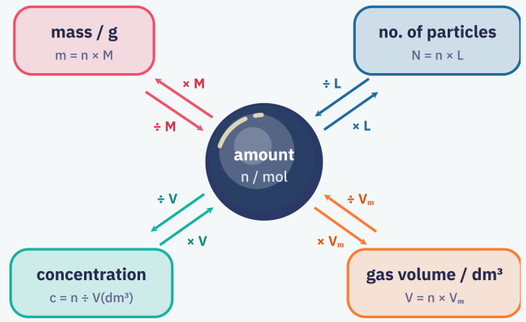
2 The Mole

DEFINITION **Mole**: the amount of substance that contains _____ (the Avogadro constant, L) elementary entities. The entity must be stated: atoms, molecules, ions or electrons.

DEFINITION **Molar mass, M**: the mass of one mole of the specified substance (g mol^{-1}), numerically equal to A_r or M_r.

DEFINITION **Molar gas volume, V_m**: the volume occupied by one mole of a gas under given conditions of temperature and pressure: _____ $dm^3\ mol^{-1}$ at r.t.p. (20 °C, 1 atm), _____ $dm^3\ mol^{-1}$ at s.t.p. (0 °C, 1 bar).

DEFINITION **Concentration**: the amount or mass of solute in a given volume of solution (mol dm^{-3} or g dm^{-3}).



Mole ratio: in 1 mol of A_mB_n there are m mol of A and n mol of B, e.g. 1 mol $CaCl_2$ has 2 mol Cl^-.

NOTE "1 mole of oxygen" is ambiguous: say 1 mol of O atoms or 1 mol of O_2 molecules. 1 dm^3 = 1000 cm^3.

2.1 Particles and gas volume

WORKED EXAMPLE A cylinder releases 16 dm^3 of NH_3 gas at s.t.p. How many NH_3 molecules is this?

$$n_{NH_3} = \frac{16}{22.7} = 0.7048\ \text{mol} \rightarrow N = 0.7048 \times 6.02 \times 10^{23} = ____$$

2.2 Preparing a solution

Weigh the solute, dissolve it in a little water, then make up to the mark in a volumetric flask and invert to mix.

WORKED EXAMPLE A technician makes 500 cm^3 of H_2SO_4 solution with $[H^+] = 2.00$ mol dm^{-3}. What mass of H_2SO_4 is needed?

$n_{H^+} = 2.00 \times 0.500 = 1.000$ mol. Each H_2SO_4 gives $2H^+$

$n_{H_2SO_4} = 0.5000$ mol → mass = 0.5000 × 98.1 = _____

3 Empirical and Molecular Formulae

DEFINITION **Empirical formula**: the formula showing the _____ of the atoms of each element in the substance.

DEFINITION **Molecular formula**: the formula showing the _____ of atoms of the different elements in one molecule of the compound, a multiple of the empirical formula.

ANSWER STRUCTURE **Finding the formula: fill in the table**

	element 1	element 2	element 3
Mass / g (or %)	given	given	given
Amount / mol	mass ÷ A_r	mass ÷ A_r	mass ÷ A_r
Ratio (÷ smallest)	whole number	whole number	whole number

Ratio → empirical formula → molecular formula = empirical formula × $\frac{M_r}{\text{empirical formula mass}}$

3.1 From % by mass

WORKED EXAMPLE Compound X (M_r = 114) is 63.2% C, 12.2% H and 24.6% N by mass. Find its molecular formula.

	C	H	N
Mass / g	63.2	12.2	24.6
Amount / mol	5.267	12.20	1.757
Ratio (÷ 1.757)	____	____	____

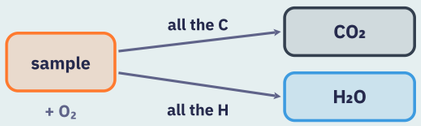
Take 100 g of X, so each percentage becomes a mass in grams.

Empirical C_3H_7N (M = 57.0), 114 ÷ 57.0 = 2 ∴ _____

3.2 From combustion data

WORKED EXAMPLE A sugar, mannose (C, H and O only, M_r 180), is burnt completely in oxygen. 2.36 g of it gives 3.46 g of CO_2 and 1.42 g of H_2O. Find its molecular formula.

1. All the C in the CO_2 came from the sample (12 g of C in every 44 g of CO_2): mass of C = $\frac{12}{44}$ × 3.46 = 0.9436 g
2. All the H in the H_2O came from the sample (2 g of H in every 18 g of H_2O): mass of H = $\frac{2}{18}$ × 1.42 = 0.1578 g
3. O cannot be found this way, as some O in CO_2 and H_2O came from the O_2 burnt: mass of O = 2.36 − 0.9436 − 0.1578 = 1.259 g



	C	H	O
Mass / g	0.9436	0.1578	1.259
Amount / mol	0.07863	0.1578	0.07869
Ratio (÷ 0.07863)	____	____	____

Empirical CH_2O (M = 30), 180 ÷ 30 = 6 ∴ _____

4 Stoichiometry, Limiting Reagent and Yield

- A balanced equation gives the _____ of reactants and products: $2H_2 + O_2 \rightarrow 2H_2O$ means 2 mol H_2 ≡ 1 mol O_2 ≡ 2 mol H_2O
- The **limiting reagent** is used up first, it fixes the maximum amount of product (the **theoretical yield**)

recipe	**2 bread**	+	**1 cheese**	→	**1 sandwich**
we have	10 bread		4 cheese		4 sandwiches
	2 left over: in **excess**		all used: **limiting**		the theoretical yield

$$\text{percentage yield} = \frac{\text{actual mass (or amount) of product}}{\text{theoretical mass (or amount) of product}} \times 100\%$$

Actual yield is lower: incomplete reaction, side reactions and losses during purification.

4.1 Limiting reagent and yield

WORKED EXAMPLE A student precipitates $PbCl_2$ by mixing 100 cm^3 of 0.0600 mol dm^{-3} $Pb(NO_3)_2$ with 200 cm^3 of 0.0500 mol dm^{-3} NaCl, and collects 1.1558 g of $PbCl_2$. Find the limiting reagent and the percentage yield.

1. $Pb(NO_3)_2 + 2NaCl \rightarrow PbCl_2 + 2NaNO_3$
2. $n(Pb(NO_3)_2) = 6.000 \times 10^{-3}$ mol, $n_{NaCl} = 1.000 \times 10^{-2}$ mol
3. 6.000×10^{-3} mol $Pb(NO_3)_2$ needs 1.200×10^{-2} mol NaCl → _____ is limiting
4. $n_{PbCl_2} = ½ \times 1.000 \times 10^{-2} = 5.000 \times 10^{-3}$ mol
5. maximum mass = $5.000 \times 10^{-3} \times 278.1$ = _____
6. % yield = $\frac{1.1558}{1.391} \times 100$ = _____

4.2 Reacting mass to gas volume

WORKED EXAMPLE A 0.230 g piece of sodium is dropped into excess water. Find the volume of H_2 produced at s.t.p.

1. $2Na + 2H_2O \rightarrow 2NaOH + H_2$. Water is in excess, so Na is limiting
2. $n_{Na} = 0.230 \div 23.0 = 1.000 \times 10^{-2}$ mol → $n_{H_2} = 5.000 \times 10^{-3}$ mol
3. $V(H_2) = 5.000 \times 10^{-3} \times 22.7$ = _____

EXAM TIP Give intermediate answers to 4 s.f. and final answers to 3 s.f. unless told otherwise, A_r and M_r to 1 d.p.

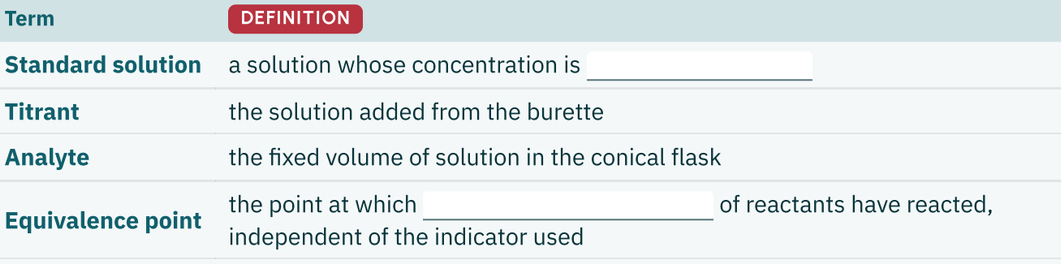
5 Gas Volumes and Combustion

5.1 Gas volumes

- **Avogadro's law**: at the same temperature and pressure, equal volumes of all gases contain the same number of particles → for gases, _____
- Only gases count: water at r.t.p. is a **liquid**, so its volume is ignored

5.2 Combustion of hydrocarbons

$$C_xH_y(g) + \left(x + \frac{y}{4}\right)O_2(g) \rightarrow xCO_2(g) + \frac{y}{2}H_2O(l)$$



$NaOH(aq)$ absorbs CO_2: $CO_2(g) + 2NaOH(aq) \rightarrow Na_2CO_3(aq) + H_2O(l)$

WORKED EXAMPLE A chemist explodes 10 cm³ of a gaseous hydrocarbon C_xH_y with 110 cm³ of O_2 (excess). After cooling to r.t.p. the volume is 85 cm³, falling to 35 cm³ after passing through excess NaOH(aq). Find the formula of the hydrocarbon.

1. CO_2 is absorbed by NaOH: $V(CO_2) = 85 - 35 = 50$ cm³ $= 10x$ → _____
2. Contraction on combustion = (10 + 110) − 85 = 35 cm³
3. Contraction = $V(C_xH_y) + V(O_2 \text{ reacted}) - V(CO_2) = 10 + 10\left(x + \frac{y}{4}\right) - 10x = 10 + \frac{10y}{4} = 35$ → _____
4. ∴ the hydrocarbon is _____

6 Volumetric Analysis

Term	DEFINITION
Standard solution	a solution whose concentration is _____
Titrant	the solution added from the burette
Analyte	the fixed volume of solution in the conical flask
Equivalence point	the point at which _____ of reactants have reacted, independent of the indicator used
End-point	the point when the indicator _____

6 Volumetric Analysis (continued)

6.1 Choosing an acid–base indicator

Titration	Indicator	Colour, acid → alkali
strong acid–strong base	methyl orange or phenolphthalein	either change below
strong acid–weak base	methyl orange (pH 3–5)	**red** → **yellow**, orange at the end-point
weak acid–strong base	phenolphthalein (pH 8–10)	colourless → **pale pink** at the end-point
weak acid–weak base	no suitable indicator	

A suitable indicator has a **distinct colour change** at the equivalence point.

6.2 Dilution

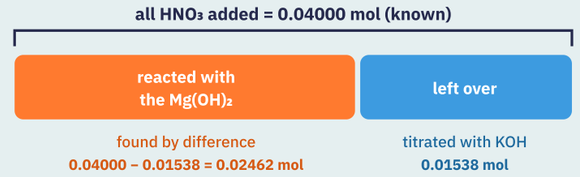
$c_1V_1 = c_2V_2$ Amount of solute withdrawn = amount in the diluted solution.

6.3 Back-titration

When A cannot be titrated directly (e.g. insoluble $CaCO_3$ or $Mg(OH)_2$)

1. Add a known **excess** of B, so all of A reacts
2. Titrate the **unreacted** B that is left over
3. $n_{B \text{ reacted with A}} = n_{B \text{ added}} - n_{B \text{ left over}}$, then n_A from the equation

WORKED EXAMPLE An antacid tablet contains impure $Mg(OH)_2$, which is insoluble, so it cannot be titrated directly. 1.00 g of it is added to 100 cm³ of 0.400 mol dm⁻³ HNO_3 (excess). The leftover HNO_3 needs 20.50 cm³ of 0.750 mol dm⁻³ KOH. Find the percentage of $Mg(OH)_2$.



$Mg(OH)_2 + 2HNO_3 \rightarrow Mg(NO_3)_2 + 2H_2O$

1. $n_{HNO_3 \text{ added}} = 0.400 \times 0.100 = 0.04000$ mol
2. $n_{HNO_3 \text{ left over}} = n_{KOH} = 0.750 \times 0.02050 = 0.01538$ mol
3. $n(HNO_3 \text{ reacted with } Mg(OH)_2) = 0.04000 - 0.01538 = 0.02462$ mol → $n(Mg(OH)_2) = ½ \times 0.02462 = 0.01231$ mol
4. mass = 0.01231 × 58.3 = 0.7177 g → _____ $Mg(OH)_2$

7 Redox Reactions

Oxidation	Reduction
gain of oxygen	loss of oxygen
loss of hydrogen	gain of hydrogen
_____	_____
increase in oxidation number	decrease in oxidation number



Zn
loses 2e⁻: **oxidised**, the **reducing agent**

Cu²⁺
gains 2e⁻: **reduced**, the **oxidising agent**

DEFINITION **Oxidising agent**: oxidises another species and is itself _____.

DEFINITION **Reducing agent**: reduces another species and is itself _____.

DEFINITION **Disproportionation**: a redox reaction in which one species is simultaneously oxidised and reduced, e.g. $Cl_2 + 2OH^- \rightarrow ClO^- + Cl^- + H_2O$, with the oxidation numbers of Cl under it: Cl_2 **0** → Cl in ClO^- **+1** (oxidised) and Cl^- **−1** (reduced).

7.1 Oxidation number rules, in order of priority

Apply them from the top: when two rules disagree, the higher one wins.

1. Element: **0** (Mg, O_2, P_4)
2. Sum of O.N. = overall _____ of the species (Al^{3+}: +3)
3. Group 1: +1, Group 2: +2, Al: +3
4. F: −1
5. H: +1 (so −1 in NaH, where Na comes first)
6. O: −2 (so −1 in H_2O_2 and +2 in OF_2)
7. Otherwise, the more electronegative atom takes the negative O.N.

WORKED EXAMPLE Find the oxidation number of Mn in MnO_4^-.

$x + 4(-2) = -1 \rightarrow x =$ _____

Charge is written "2+", oxidation number is written "+2".

7.2 Balancing redox equations: MnO_4^- + I^- in acid

AOHC Half-equation method

A	Balance **atoms** other than O and H	$MnO_4^- \rightarrow Mn^{2+}$ $2I^- \rightarrow I_2$
O	Balance **O** with H_2O	$MnO_4^- \rightarrow Mn^{2+} + 4H_2O$
H	Balance **H** with H^+	$MnO_4^- + 8H^+ \rightarrow Mn^{2+} + 4H_2O$
C	Balance **charge** with e^-	$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$ $2I^- \rightarrow I_2 + 2e^-$
Add	Match the e^- (× 2 and × 5), add, cancel	_____

ACOH Oxidation number method

A	Balance **atoms** other than O and H	$MnO_4^- + 2I^- \rightarrow Mn^{2+} + I_2$
C	Balance the **change** in O.N.: Mn −5 each, I +1 each	$2MnO_4^- + 10I^- \rightarrow 2Mn^{2+} + 5I_2$
O	Balance **O** with H_2O	$2MnO_4^- + 10I^- \rightarrow 2Mn^{2+} + 5I_2 + 8H_2O$
H	Balance **H** with H^+	_____

NOTE In **alkaline** solution MnO_4^- is reduced to MnO_2. Balance with H^+ first, then add the same number of OH^- to both sides to turn H^+ into H_2O.

8 Redox Titrations

Titration	Half-equation of oxidising agent	Indicator	End-point colour
$KMnO_4$ (burette) with Fe^{2+} or H_2O_2	$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$	none (self-indicating)	colourless → _____
	↳ Acidify with **H_2SO_4**, not HCl (Cl^- is oxidised to Cl_2) or HNO_3 (itself an oxidising agent)		
$K_2Cr_2O_7$ with Fe^{2+}	$Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$	N-phenylanthranilic acid	green → red-violet
$S_2O_3^{2-}$ (burette) with I_2	$I_2 + 2S_2O_3^{2-} \rightarrow 2I^- + S_4O_6^{2-}$	starch, near the end	dark blue → _____
	↳ **Iodometric titration**: the oxidising agent is added to _____, so it reacts completely and the I_2 stays dissolved as I_3^-. The I_2 released is then titrated with $S_2O_3^{2-}$: the titre gives the amount of I_2, and so the original amount of oxidising agent (e.g. IO_3^-)		
	↳ Add starch only near the end-point (pale yellow): at high $[I_2]$ the starch–iodine complex precipitates		

8.1 Iodometric titration

WORKED EXAMPLE To standardise a sodium thiosulfate solution, 0.875 g of KIO_3 is made up to 250 cm³. 25.0 cm³ of this is added to excess KI and acid, and the I_2 formed needs 22.70 cm³ of $Na_2S_2O_3$. Find $[S_2O_3^{2-}]$.

1. $IO_3^- + 5I^- + 6H^+ \rightarrow 3I_2 + 3H_2O$, $I_2 + 2S_2O_3^{2-} \rightarrow 2I^- + S_4O_6^{2-}$
2. $n_{IO_3^-}$ in 25.0 cm³ = $\frac{0.875}{214.0} \times \frac{25.0}{250} = 4.089 \times 10^{-4}$ mol
3. $IO_3^- \equiv 3I_2 \equiv$ _____ → $n_{S_2O_3^{2-}} = 2.453 \times 10^{-3}$ mol
4. $[S_2O_3^{2-}] = 2.453 \times 10^{-3} \div 0.02270 =$ _____

8.2 Finding a new oxidation number

electrons lost = electrons gained

30.0 cm³ of 0.0500 mol dm⁻³ Fe^{2+} reacts exactly with 15.0 cm³ of 0.0250 mol dm⁻³ MnO_4^-, and Fe ends as +3. Find the final oxidation number of Mn.

1. $n_{Fe^{2+}} = 1.500 \times 10^{-3}$ mol, each losing 1 e^- → 1.500×10^{-3} mol e^-
2. $n_{MnO_4^-} = 3.750 \times 10^{-4}$ mol → each gains $\frac{1.500 \times 10^{-3}}{3.750 \times 10^{-4}} =$ _____ e^-
3. Mn: +7 − 4 = _____

8.3 Common oxidising and reducing agents

Oxidising agent	Medium	Product	Reducing agent	Medium	Product
MnO_4^-	acidic	Mn^{2+}	Fe^{2+}	acidic or neutral	Fe^{3+}
MnO_4^-	neutral or alkaline	MnO_2	I^-	any	I_2
$Cr_2O_7^{2-}$	acidic	Cr^{3+}	$S_2O_3^{2-}$	neutral	$S_4O_6^{2-}$ with I_2
I_2	neutral or acidic	I^-	$C_2O_4^{2-}$	acidic	CO_2
IO_3^-	acidic	I_2	H_2O_2	acidic	O_2
H_2O_2	acidic	H_2O			

H_2O_2 can act as an oxidising agent or as a reducing agent. With a stronger oxidising agent such as Cl_2, $S_2O_3^{2-}$ gives SO_4^{2-}.