Topic 5: Formulae, Equations and Amounts of Substance

Edexcel A Level Chemistry • Paper 1 (9CH0/01)

5.1 Know that the mole (mol) is the unit for amount of a substance.

The mole is the SI unit for amount of substance. One mole contains exactly 6.02×10^{23} particles (Avogadro's constant, L). The particles can be atoms, molecules, ions, electrons, or formula units – always specify the type.

5.2 Be able to use the Avogadro constant, L, (6.02×10^{23} mol^{-1}) in calculations.

Number of particles = moles × L

Moles = number of particles / L

E.g. How many molecules in 0.5 mol of H_2O? $0.5 \times 6.02 \times 10^{23} = 3.01 \times 10^{23}$ molecules.

5.3 Know that the molar mass of a substance is the mass per mole of the substance in g mol^{-1}.

Molar mass (M) has units of g mol^{-1} and is numerically equal to the relative atomic/molecular/formula mass. E.g. $M(H_2O)$ = 18.0 g mol^{-1}.

Moles = mass / molar mass (n = m/M).

5.4 Know what is meant by the terms 'empirical formula' and 'molecular formula'.

Empirical formula: the simplest whole number ratio of atoms of each element in a compound (e.g. CH_2O).

Molecular formula: the actual number of atoms of each element in one molecule (e.g. $C_6H_{12}O_6$ = glucose). Molecular formula is always a whole number multiple of the empirical formula.

5.5 Be able to use experimental data to calculate: (i) empirical formulae, (ii) molecular formulae including the use of pV = nRT for gases and volatile liquids.

(i) Empirical formula method: divide each mass (or %) by the Ar to get moles; divide all by the smallest to get the ratio; scale up to whole numbers if needed.

(ii) Molecular formula: find Mr (e.g. from mass spec or pV = nRT), then divide Mr by the empirical formula mass to find the multiplier.

Ideal gas equation: pV = nRT where p = pressure (Pa), V = volume (m^3), n = moles, R = 8.314 J K^{-1} mol^{-1}, T = temperature (K). Convert units carefully.

© Chemistry with Chloe | chemistrywithchloe.com
For personal educational use only. Not for redistribution.

5.6

Be able to write balanced full and ionic equations, including state symbols, for chemical reactions.

Full equations: show all reactants and products with correct formulae and state symbols: (s) solid, (l) liquid, (g) gas, (aq) aqueous.

Ionic equations: only show the species that change during the reaction. Remove spectator ions (those unchanged on both sides). Always check charges balance as well as atoms.

5.7

Be able to calculate amounts of substances (in mol) in reactions involving mass, volume of gas, volume of solution and concentration.

Key formulae:

- $n = m/M$ (mass and molar mass)
- $n = V/24{,}000$ (gas volume in cm^3 at RTP) or $n = V/24$ (dm^3)
- $n = c \times V$ (concentration in mol dm^{-3} × volume in dm^3)
- $pV = nRT$ (ideal gas equation)

Always convert units: 1 dm^3 = 1000 cm^3.

5.8

Be able to calculate reacting masses from chemical equations, and vice versa.

Method: write a balanced equation → calculate moles of the known substance → use the mole ratio from the equation to find moles of the unknown → convert to mass using $n = m/M$. Always identify the limiting reagent if amounts of both reactants are given.

5.9

Be able to calculate reacting volumes of gases from chemical equations using the concepts of amount of substance.

Use $pV = nRT$ to calculate moles from gas volumes at non-standard conditions. Then use the mole ratio to find the required amount and convert back to volume. Remember to convert T to Kelvin and V to m^3 (or use consistent units).

5.10

Be able to calculate reacting volumes of gases from chemical equations using the concepts of molar volume of gases.

At room temperature and pressure (RTP), 1 mole of any gas occupies approximately 24 dm^3 (24,000 cm^3). Equal volumes of gases at the same T and P contain equal numbers of moles (Avogadro's law). Gas volume ratios = mole ratios from the balanced equation.

5.11

Be able to calculate solution concentrations, in mol dm^{-3} and g dm^{-3}, including simple acid-base titrations.

Concentration (mol dm^{-3}): $c = n/V$ (moles ÷ volume in dm^3)

Concentration (g dm^{-3}): = $c \times M$ (mol dm^{-3} × molar mass)

Titration calculations: use concordant titres (within 0.10 cm^3). Calculate moles of known → use mole ratio → calculate concentration of unknown.

Indicators: phenolphthalein (colourless in acid, pink in alkali), methyl orange (red in acid, yellow in alkali).

© Chemistry with Chloe | chemistrywithchloe.com
For personal educational use only. Not for redistribution.

5.12 Be able to: (i) calculate measurement uncertainties and measurement errors, (ii) comment on sources of error.

(i) Percentage uncertainty = (uncertainty / measurement) × 100%

Burette: ±0.05 cm^3 per reading (±0.10 cm^3 for a titre = 2 readings). Balance: ±0.001 g (or ±0.01 g). Volumetric flask: ±0.1 cm^3 (250 cm^3). Pipette: ±0.06 cm^3 (25 cm^3).

(ii) Sources of error: systematic (affect accuracy, e.g. zero error on balance, indicator colour change not at true end point) and random (affect precision, e.g. reading meniscus, judging colour change).

5.13 Understand how to minimise the percentage error and percentage uncertainty in experiments involving measurements.

Minimise percentage uncertainty by making measurements as large as possible. E.g. use a larger mass on the balance, use a burette rather than a measuring cylinder, ensure titre volumes are not too small. Repeat measurements to reduce random error and improve precision. Use more precise equipment (e.g. volumetric pipette instead of measuring cylinder).

5.14 Be able to calculate percentage yields and percentage atom economies using chemical equations and experimental results.

Percentage yield = (actual yield / theoretical yield) × 100%

Yield < 100% due to: incomplete reaction, side reactions, losses during transfers/purification.

Atom economy = (molar mass of desired product / sum of molar masses of ALL products) × 100%

High atom economy = less waste. Addition reactions have 100% atom economy. Substitution/elimination reactions have lower atom economy.

5.15 Be able to relate ionic and full equations to observations from simple test tube reactions.

You must be able to write equations AND describe what you would see:

(i) Displacement: e.g. $Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$. Observation: grey zinc dissolves, blue solution fades, brown copper deposited.

(ii) Acids: e.g. $Mg + 2HCl \rightarrow MgCl_2 + H_2$. Observation: effervescence, metal dissolves.

(iii) Precipitation: e.g. $Pb^{2+}(aq) + 2I^-(aq) \rightarrow PbI_2(s)$. Observation: yellow precipitate forms.

5.16 Understand risks and hazards in practical procedures and suggest appropriate precautions.

Hazard = something that can cause harm. Risk = the likelihood of harm occurring.

Precautions include: wearing safety goggles (corrosive/irritant substances), working in a fume cupboard (toxic/volatile gases), using small quantities, keeping away from naked flames (flammable substances), washing hands after handling chemicals. Always link the precaution to the specific hazard.

Measure the Molar Volume of a Gas

© Chemistry with Chloe | chemistrywithchloe.com
For personal educational use only. Not for redistribution.

Prepare a Standard Solution from a Solid Acid and Use It to Find the Concentration of NaOH

Find the Concentration of a Solution of Hydrochloric Acid

ABOUT THESE NOTES

Notes written by Chloe Archard (MSc Biochemistry, Oxford-trained chemist, 20+ years teaching experience at top UK independent schools). For 1-1 tuition, revision courses, and more free resources: chemistrywithchloe.com

© Chemistry with Chloe | chemistrywithchloe.com
For personal educational use only. Not for redistribution.