CHEMISTRY
TOPIC 1

Complete Typed Study Notes
Atoms and Ions | Chemical Equations | Mole Concept
Stoichiometry | Formulae | Gas Laws | Concentration

About this document
These notes have been typed and organised from the supplied 12-page handwritten Topic 1 document. The original subject matter, formulae, worked examples and revision points have been arranged into readable sections.

For Chemistry learning and revision

1. Atoms, Elements, Compounds and Molecules

Atom: The smallest particle of an element that retains the chemical identity of that element. An atom contains protons, neutrons and electrons.

Element: A substance made up of only one type of atom. Examples include hydrogen (H), oxygen (O), iron (Fe) and copper (Cu).

Compound: A substance formed when atoms of two or more different elements are chemically bonded together in a fixed ratio. Examples include water (H_2O), carbon dioxide (CO_2) and methane (CH_4).

Molecule: Two or more atoms chemically bonded together. The atoms may belong to the same element or to different elements.

Type	Description	Examples
Elemental molecule	Contains atoms of the same element.	H_2, N_2, O_2, Cl_2, P_4, S_8
Compound molecule	Contains atoms of different elements.	H_2O, CO_2, CH_4, NH_3

Important distinction
Not all compounds are made of molecules. Ionic compounds such as sodium chloride form giant ionic lattices.

Ions

Ion: A charged particle formed when an atom or a group of atoms gains or loses electrons.

Ion type	How it forms	Examples
Cation (positive)	Formed when electrons are lost.	Na^+, Mg^{2+}, Al^{3+}
Anion (negative)	Formed when electrons are gained.	Cl^-, O^{2-}, N^{3-}

A positive ion has fewer electrons than protons. A negative ion has more electrons than protons.

Monatomic ion: A single atom carrying a charge, e.g. Na^+ or Cl^-. **Polyatomic ion:** A group of covalently bonded atoms carrying an overall charge, e.g. NH_4^+, NO_3^-, SO_4^{2-} and CO_3^{2-}.

Common ions

Positive ions	Negative ions
H^+ hydrogen; Na^+ sodium; K^+ potassium	OH^- hydroxide; Cl^- chloride; Br^- bromide; I^- iodide
NH_4^+ ammonium; Ag^+ silver	NO_3^- nitrate; HCO_3^- hydrogencarbonate
Mg^{2+} magnesium; Ca^{2+} calcium; Cu^{2+} copper(II)	CO_3^{2-} carbonate; SO_4^{2-} sulfate; SO_3^{2-} sulfite; O^{2-} oxide
Fe^{2+} iron(II); Fe^{3+} iron(III)	PO_4^{3-} phosphate; N^{3-} nitride

2. Solubility Rules and Ionic Reactions

Solubility rules help predict whether an ionic compound dissolves in water or forms an insoluble solid (a precipitate).

Compound type	Solubility rule
Nitrates	All nitrates are soluble.
Group 1 and ammonium salts	All Group 1 salts and ammonium salts are soluble.
Chlorides, bromides and iodides	Usually soluble, except those of silver and lead(II).
Sulfates	Usually soluble, except barium sulfate and lead(II) sulfate. Calcium sulfate is sparingly soluble.
Carbonates	Usually insoluble, except Group 1 and ammonium carbonates.
Hydroxides	Usually insoluble, except Group 1 hydroxides. Calcium hydroxide is slightly soluble; barium hydroxide is soluble.

Precipitation reactions

A precipitation reaction occurs when two aqueous ionic solutions are mixed and an insoluble solid forms.

$$BaCl_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2NaCl(aq)$$

Barium sulfate is a white precipitate. The net ionic equation is:

$$Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$$

Tests for halide ions

Acidify the sample with dilute nitric acid, then add aqueous silver nitrate.

Halide ion	Observation	Precipitate
Chloride, Cl^-	White precipitate	AgCl
Bromide, Br^-	Cream precipitate	AgBr
Iodide, I^-	Yellow precipitate	AgI

Nitric acid removes interfering carbonate ions. Hydrochloric acid must not be used because it introduces chloride ions.

Acids, bases and neutralisation

$$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$$

Acids produce $H^+(aq)$ ions in aqueous solution. Alkalis are soluble bases that produce $OH^-(aq)$ ions. A monoprotic acid donates one H^+ per molecule; a triprotic acid can donate three. Phosphoric acid, H_3PO_4, is triprotic.

3. Chemical Equations and Reaction Types

A chemical equation represents a chemical reaction using formulae. A balanced equation contains the same number of atoms of each element on both sides.

State symbols

Symbol	Meaning
(s)	Solid
(l)	Liquid
(g)	Gas
(aq)	Aqueous - dissolved in water

Balancing equations

1. Write the correct formulae of all reactants and products.
2. Count the atoms of each element on both sides.
3. Add whole-number coefficients in front of formulae until the number of each type of atom is equal.
4. Do not change the small numbers (subscripts) within formulae.
5. Add state symbols when required.

Common reaction patterns

Metal + acid → salt + hydrogen

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$

Acid + carbonate → salt + water + carbon dioxide

$$2HNO_3(aq) + CaCO_3(s) \rightarrow Ca(NO_3)_2(aq) + H_2O(l) + CO_2(g)$$

Metal displacement: A more reactive metal displaces a less reactive metal from its compound.

$$Mg(s) + CuSO_4(aq) \rightarrow MgSO_4(aq) + Cu(s)$$

$$2Al(s) + Fe_2O_3(s) \rightarrow Al_2O_3(s) + 2Fe(l)$$

In the thermite reaction, aluminium displaces iron from iron(III) oxide.

Halogen displacement: A more reactive halogen displaces a less reactive halide from solution. Reactivity decreases down Group 17:

$$F_2 > Cl_2 > Br_2 > I_2$$

$$Cl_2(aq) + 2NaBr(aq) \rightarrow 2NaCl(aq) + Br_2(aq)$$

Precipitation:

$$BaCl_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2NaCl(aq)$$

Net ionic equations

A net ionic equation includes only the particles that take part in the reaction. Ions that remain unchanged in solution are spectator ions.

$$Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$$

$$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$$

4. Relative Masses, Isotopes and the Mole

Relative atomic mass (A_r)

Relative atomic mass is the weighted mean mass of an atom of an element compared with one-twelfth of the mass of a carbon-12 atom. It has no units.

Relative molecular mass (M_r) and formula mass

Relative molecular mass is the sum of the relative atomic masses of all atoms in a molecule. For ionic compounds, the equivalent calculation gives the relative formula mass.

$$M_r = \text{sum of the } A_r \text{ values of all atoms in the formula}$$

Example: CO_2

$$M_r = 12.0 + (2 \times 16.0) = 44.0$$

Example: $Ca(NO_3)_2$

$$M_r = 40.1 + 2(14.0 + 3 \times 16.0) = 164.1$$

Isotopes and mass spectrum

Isotopes are atoms of the same element with the same number of protons but different numbers of neutrons. A mass spectrum shows the mass-to-charge ratio and relative abundance of ions.

$$A_r = \text{sum of (isotopic mass} \times \text{percentage abundance)} \div 100$$

Example: Isotopes with masses 24 and 25 have abundances of 60% and 40%.

$$A_r = [(24 \times 60) + (25 \times 40)] \div 100 = 24.4$$

The mole and Avogadro constant

One mole contains 6.022×10^{23} specified particles. This is the Avogadro constant, L.

$$L = 6.022 \times 10^{23} \text{ mol}^{-1}$$

The mass of one mole of a substance in grams is numerically equal to its relative atomic, molecular or formula mass.

Term	Meaning
Atomic mass unit (u)	A unit for the mass of individual atoms or particles. One carbon-12 atom has a mass of exactly 12 u.
Mole (mol)	An amount of substance containing 6.022×10^{23} specified particles.

5. Mole Calculations and Interconversions

Mass and moles

$$n = m \div M \quad m = n \times M \quad M = m \div n$$

n = amount in mol; m = mass in g; M = molar mass in g mol^{-1}.

Moles and particles

$$\text{Number of particles} = n \times L$$

$$n = \text{number of particles} \div L$$

Moles in gases

At room temperature and pressure (RTP), the molar gas volume is approximately 24.0 dm^3 mol^{-1}.

$$n = \text{gas volume } (dm^3) \div 24.0$$

Always use the molar gas volume specified by the question if a different condition or value is given.

Useful unit conversions

Quantity	Conversion
Volume	1 dm^3 = 1000 cm^3
Volume	1 cm^3 = 1 mL
Mass	1 kg = 1000 g
Temperature	T(K) = temperature (°C) + 273

Mole-map method for reacting-mass calculations

For a mass-to-mass calculation, convert the given mass into moles, use the balanced equation to find the mole ratio, then convert the required moles into mass.

1. Write a balanced chemical equation.
2. Calculate moles of the known substance using n = m ÷ M.
3. Use the coefficients in the equation to obtain the mole ratio.
4. Calculate moles of the required substance.
5. Convert the answer into the required quantity, such as mass or gas volume.

Example: $2Mg + O_2 \rightarrow 2MgO$. If 4.8 g of magnesium reacts completely:

$$n(Mg) = 4.8 \div 24.3 = 0.1975 \text{ mol}$$

$$Mg : MgO = 2 : 2, \text{ so } n(MgO) = 0.1975 \text{ mol}$$

$$m(MgO) = 0.1975 \times 40.3 = 7.96 \text{ g}$$

Limiting and excess reactants

Limiting reactant: The reactant used up first; it determines the maximum amount of product. **Excess reactant:** The reactant left over.

1. Calculate the moles of each reactant.
2. Compare the mole amounts with the coefficients in the balanced equation.
3. Divide each available mole amount by its coefficient; the smaller result identifies the limiting reactant.
4. Use the limiting reactant to calculate the theoretical yield.

Important
The reactant with the smallest mass or number of moles is not automatically the limiting reactant. Compare the available moles using the balanced equation.

6. Reacting Masses, Percentage Yield and Atom Economy

Reacting masses

A balanced equation gives the mole ratio in which reactants combine and products form. Use the mole-map method in Section 5.

Percentage yield

Theoretical yield is the maximum amount predicted by the balanced equation. Actual yield is the amount obtained experimentally.

$$\% \text{ yield} = (\text{actual yield} \div \text{theoretical yield}) \times 100$$

A yield below 100% may result from incomplete reaction, side reactions or loss of product during separation and purification.

Atom economy

Atom economy measures the proportion of the atoms in the reactants that become part of the desired product.

$$\% \text{ atom economy} = [M_r \text{ of desired product} \div \text{total } M_r \text{ of all products}] \times 100$$

A high atom economy means less waste is produced.

Example: $CaCO_3 \rightarrow CaO + CO_2$. If CaO is the desired product:

$$\% \text{ atom economy} = [56.1 \div (56.1 + 44.0)] \times 100 = 56.0\%$$

Addition, substitution and elimination reactions

Addition: Two reactants combine to form one product. In addition polymerisation, many alkene monomers join to form a polymer. The C=C double bond opens and forms new C-C single bonds.

$$n\ CH_2{=}CH_2 \rightarrow [\text{-}CH_2\text{-}CH_2\text{-}]_n$$

Substitution: An atom or group in a molecule is replaced by another atom or group.

Elimination: A small molecule is removed from a larger molecule, often forming a double bond.

Combustion of hydrocarbons: Complete combustion produces carbon dioxide and water.

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

7. Empirical and Molecular Formulae

Empirical formula: The simplest whole-number ratio of atoms of each element in a compound.

Molecular formula: The actual number of atoms of each element in one molecule.

Finding an empirical formula from masses or percentages

1. Write down the mass or percentage of each element.
2. Divide each value by the relative atomic mass (A_r) to calculate the amount in moles.
3. Divide all mole values by the smallest value.
4. If necessary, multiply all ratios by the same number to obtain whole numbers.

Ratio obtained	Multiply by	Whole-number ratio
1 : 1.5	2	2 : 3
1 : 1.33	3	3 : 4
1 : 1.25	4	4 : 5
1 : 1.67	3	3 : 5
1 : 2.5	2	2 : 5

Example: A compound contains 40.0% carbon, 6.7% hydrogen and 53.3% oxygen.

Element	Percentage	Divide by A_r	Mole ratio
C	40.0	40.0 ÷ 12 = 3.33	1
H	6.7	6.7 ÷ 1 = 6.7	2
O	53.3	53.3 ÷ 16 = 3.33	1

Empirical formula = CH_2O

Finding the molecular formula

Multiplier = M_r ÷ empirical formula mass

Multiply every subscript in the empirical formula by this whole-number multiplier.

Example: Empirical formula CH_2O; molecular mass = 180.

Empirical formula mass = 12 + (2 × 1) + 16 = 30

Multiplier = 180 ÷ 30 = 6

Molecular formula = $C_6H_{12}O_6$

8. Combustion Analysis

In combustion analysis, a known mass of a compound is burned completely in excess oxygen. Carbon forms carbon dioxide and hydrogen forms water.

Carbon in compound → CO_2 Hydrogen in compound → H_2O

Calculating carbon and hydrogen

- Moles of CO_2 = mass of CO_2 ÷ 44.0. Each mole of CO_2 contains one mole of carbon atoms.
- Moles of H_2O = mass of H_2O ÷ 18.0. Each mole of water contains two moles of hydrogen atoms.
- Moles of hydrogen atoms = 2 × moles of H_2O.
- If oxygen is present in the original compound, calculate its mass by subtracting the masses of carbon and hydrogen from the original sample mass.
- Convert the mass of each element into moles and find the simplest whole-number ratio.

n(C) = mass of CO_2 ÷ 44.0

n(H) = 2 × [mass of H_2O ÷ 18.0]

mass of O = sample mass - mass of C - mass of H

Worked example: A 0.300 g sample produces 0.440 g CO_2 and 0.180 g H_2O.

n(CO_2) = 0.440 ÷ 44.0 = 0.0100 mol; therefore n(C) = 0.0100 mol

mass of C = 0.0100 × 12.0 = 0.120 g

n(H_2O) = 0.180 ÷ 18.0 = 0.0100 mol; therefore n(H) = 0.0200 mol

mass of H = 0.0200 × 1.0 = 0.0200 g

mass of O = 0.300 - 0.120 - 0.020 = 0.160 g

n(O) = 0.160 ÷ 16.0 = 0.0100 mol

C : H : O = 0.0100 : 0.0200 : 0.0100 = 1 : 2 : 1

Empirical formula = CH_2O

Common error
Remember to multiply the moles of water by 2 to obtain the moles of hydrogen atoms.

9. Gas Laws and the Ideal Gas Equation

The gas laws describe the relationships between pressure, volume, temperature and amount of gas.

Gas law	Relationship	Conditions
Boyle's law	$V \propto 1/P$	Temperature and amount of gas are constant.
Charles's law	$V \propto T$	Pressure and amount of gas are constant.
Avogadro's law	$V \propto n$	Temperature and pressure are constant.

Temperature must be in kelvin:

$$\mathbf{T(K) = \text{temperature } (°C) + 273}$$

Ideal gas equation

$$\mathbf{PV = nRT}$$

- P = pressure in pascals (Pa)
- V = volume in cubic metres (m^3)
- n = amount in moles (mol)
- R = 8.31 J mol^{-1} K^{-1}
- T = temperature in kelvin (K)

Useful conversions

Quantity	Conversion
Pressure	1 kPa = 1000 Pa
Pressure	1 atm ≈ 101 325 Pa
Volume	1 dm^3 = 1 × 10^{-3} m^3
Volume	1 cm^3 = 1 × 10^{-6} m^3

Molar mass from gas data

Using PV = nRT and n = m/M:

$$\mathbf{PV = (m/M)RT}$$

$$\mathbf{M = mRT \div PV}$$

If density d = m/V:

$$\mathbf{M = dRT \div P}$$

Exam check
Convert volume into m^3, pressure into Pa and temperature into K when using R = 8.31 in SI units.

10. Concentration of Solutions

Concentration describes the amount of solute dissolved in a given volume of solution.

Molar concentration

$$\mathbf{c = n \div V \quad n = c \times V}$$

c is in mol dm^{-3}, n is in mol and V is in dm^3. Convert cm^3 to dm^3 by dividing by 1000.

Example: Calculate moles in 25.0 cm^3 of a 0.200 mol dm^{-3} solution.

$$\mathbf{V = 25.0 \div 1000 = 0.0250\ dm^3}$$

$$\mathbf{n = 0.200 \times 0.0250 = 0.00500\ mol}$$

Mass concentration

$$\text{Mass concentration (g dm}^{-3}\text{)} = \text{mass of solute (g)} \div \text{volume (dm}^{3}\text{)}$$

$$\text{Mass concentration} = c\ \text{(mol dm}^{-3}\text{)} \times M\ \text{(g mol}^{-1}\text{)}$$

Dilution

When solvent is added, the amount of solute remains constant but the concentration decreases.

$$c_1V_1 = c_2V_2$$

Use consistent volume units.

Preparing a solution from a solid

1. Calculate the required mass using n = cV and m = nM.
2. Accurately weigh the solid.
3. Dissolve it in a small volume of distilled water in a beaker.
4. Transfer the solution quantitatively into a volumetric flask, rinsing the beaker and funnel into the flask.
5. Add distilled water until the bottom of the meniscus reaches the calibration mark.
6. Stopper and invert the flask several times to mix thoroughly.

Accuracy
The final volume is the volume of the solution, not the volume of water initially added. Use a volumetric flask to prepare a solution of accurately known volume.

11. Parts per Million (ppm)

Parts per million expresses very low concentrations, commonly used for pollutants in air and water.

Mass-based ppm

$$\text{ppm} = (\text{mass of solute} \div \text{mass of solution}) \times 10^{6}$$

For a very dilute aqueous solution with density close to that of water, 1 ppm is approximately 1 mg dm^{-3} (1 mg L^{-1}).

Converting percentage into ppm

$$\text{ppm} = \text{percentage} \times 10^{4}$$

Example: A pollutant is present at 0.00004% by mass.

$$\text{ppm} = 0.00004 \times 10^{4} = 0.4\ \text{ppm}$$

Volume-based ppm

$$\text{ppm by volume} = (\text{volume of pollutant} \div \text{total volume}) \times 10^{6}$$

Example: A solution contains 0.176 g of solute in 750 g of solution.

$$\text{ppm} = (0.176 \div 750) \times 10^{6} = 235\ \text{ppm (3 s.f.)}$$

Remember
1% = 10 000 ppm. Always check whether the question gives a mass ratio or a volume ratio.

12. Gas Volumes and Combustion of Hydrocarbons

At the same temperature and pressure, equal volumes of gases contain equal numbers of molecules. Therefore, gas volumes react in the same ratio as the coefficients in a balanced equation when all gases are measured under the same conditions.

Complete combustion

Complete combustion occurs in excess oxygen and produces carbon dioxide and water.

$$C_xH_y + (x + y/4)O_2 \rightarrow xCO_2 + (y/2)H_2O$$

Example: propane

$$C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$$

Incomplete combustion

When oxygen is limited, incomplete combustion may produce carbon monoxide and/or carbon (soot), as well as water.

$$2CH_4 + 3O_2 \rightarrow 2CO + 4H_2O$$

Gas-volume example

100 cm^3 of propane is completely burned:

1 volume C_3H_8 : 5 volumes O_2 : 3 volumes CO_2

100 cm^3 propane requires 500 cm^3 oxygen and produces 300 cm^3 carbon dioxide.

If 600 cm^3 oxygen was initially present:

Oxygen left = 600 - 500 = 100 cm^3

If the product mixture is cooled and water condenses, the final gas volume is carbon dioxide plus unreacted oxygen:

Final gas volume = 300 + 100 = 400 cm^3

> **Important**
> Water is a liquid at room temperature. If the question states that the mixture is cooled, exclude water vapour when it condenses.

13. Formula Summary and Exam Checklist

Topic	Key equation / relationship
Moles from mass	$n = m \div M$
Mass	$m = n \times M$
Particles	$N = n \times L$
Moles from particles	$n = N \div L$
Gas volume at RTP	$n = V \div 24.0$ (V in dm^3)
Molar concentration	$c = n \div V$
Moles in solution	$n = cV$
Mass concentration	mass (g) ÷ volume (dm^3)
Dilution	$c_1V_1 = c_2V_2$
Ideal gas equation	$PV = nRT$
Molar mass from gas density	$M = dRT \div P$
Percentage yield	(actual ÷ theoretical) × 100
Atom economy	(M_r desired product ÷ total M_r products) × 100
Molecular formula	Multiplier = M_r ÷ empirical formula mass
Mass ppm	(mass solute ÷ mass solution) × 10^6
Percentage to ppm	ppm = percentage × 10^4

Before submitting a calculation

- Is the chemical equation balanced?
- Have I used the correct formula and molar mass?
- Are the units consistent (cm^3 or dm^3; Pa or kPa; °C or K)?

- Have I used the mole ratio from the balanced equation?
- Have I included appropriate significant figures?
- Have I written the final unit?
- Does the answer make chemical sense?

Revision advice
Practise each calculation type separately before attempting mixed questions. Identify the information given, choose the correct relationship, show every step and check your units.