Unit 2 · Stoichiometry

STUDY NOTES

WHAT NEB'S GRADE 11 EXAM ASKS

- **CDC's grid (2077):** Physical Chemistry (Units 1–8 together): 28 of 75 marks
- **The paper:** 11 MCQs × 1 · 8 short × 5 · 3 long × 8; 3 hours; 1 mark per expected point
- **CDC's model paper:** MCQ: moles of NH_3 from 0.5 mol N_2 (1); 8-mark long question: $4\,NH_3 + 5\,O_2 \longrightarrow 4\,NO + 6\,H_2O$ with 2 mol of each, mass of NO, why the limiting reactant matters, percentage yield from 50 g of water, volume of HCl at 20 °C and 750 mmHg to react with 2 mol NH_3
- **MCQ (1 mark)**
  - formula of a second oxide from % O
  - define a mole; mol of O atoms in 0.5 mol O_2
  - molecules in 88 g CO_2; mol in 1.8 mL water
  - define the limiting reactant
  - atoms in 2x g of B from x g of A
  - volume of 1 mol O_2 vs O_3
  - O_2 to burn 0.5 mol CH_4; H_2 from 4.8 g Mg at STP
  - % yield; moles of D in $4\,A + 6\,B \longrightarrow 2\,C + 3\,D$
- **Short (5 marks, in parts)**
  - state a law with an example: multiple, reciprocal, gaseous volumes; Avogadro's law and four applications
  - differences: limiting/excess reactant, conservation of mass/constant composition, experimental/theoretical yield
  - which law a set of data shows, and prove it (% of C in CO_2, CS_2; CuS, CuO, SO_3; two oxides of carbon)
  - why atomic masses are fractions; equal numbers of Na and Al atoms on a balance
  - molecules and atoms in moles of H_2, He; 22.4 L of H_2, O_2, CO_2 at STP
- **Long (8 marks)**: limiting reactant, mass and STP volume of product, excess left, theoretical, actual and % yield ($S + O_2$, $CaCO_3 + HCl$, $NaBr + H_3PO_4$, $Ca + N_2$, $Al_2S_3 + H_2O$, $KClO_3 \rightarrow O_2$); empirical formula from % or from burning (CO_2 and H_2O)

Key terms

Stoichiometry: study of the quantities of substances in chemical reactions (Greek: element + measure; Richter, 1792).

Mole (mol): amount of substance with as many particles as there are atoms in exactly 12 g of carbon-12; since 2019, exactly $6.022\,140\,76 \times 10^{23}$ particles.

Avogadro constant N_A: 6.022×10^{23} per mol.

Molar mass M: mass of 1 mol (g/mol) = atomic or molecular mass in grams; **gram atom**, **gram molecule**: 1 mol of atoms, of molecules.

Molar volume: volume of 1 mol of gas: **22.4 L at STP/NTP** (0 °C, 1 atm); 24.5 L at RTP (25 °C, 1 atm).

Vapor density (VD): density of a gas ÷ density of hydrogen at the same T, P; no unit.

Equivalent mass: mass of an element that combines with 1 g of hydrogen or 8 g of oxygen.

Limiting reactant: used up first; decides the amount of product. **Excess reactant**: left over.

Theoretical yield: most product the equation allows from the limiting reactant. **Experimental (actual) yield**: product really obtained.

Percentage yield: actual ÷ theoretical × 100%; a share, no unit, never over 100%.

Formulas

$n = \frac{m}{M} = \frac{V_{STP}}{22.4\text{ L}} = \frac{N}{N_A}$	m g, V L at STP, N particles **Use:** moles from anything
$M_r = 2 \times \text{VD}$	VD vs hydrogen **Use:** gas's molecular mass
M = mass of 22.4 L at STP	1 L of H_2 = 0.0899 g **Use:** molar mass of a gas
$\%\text{ yield} = \frac{\text{actual}}{\text{theoretical}} \times 100$	masses, moles or volumes **Use:** efficiency
$n = \frac{\text{molecular mass}}{\text{EF mass}}$	molecular formula = $(\text{EF})_n$ **Use:** from empirical formula

The laws of chemical combination

Dalton's postulates (1808)

1. matter: tiny indivisible atoms
2. one element's atoms alike in mass and properties
3. atoms combine in simple whole-number ratios
4. a reaction only separates, rearranges and joins atoms

Conservation of mass (Lavoisier; indestructibility of matter): total mass of products = total mass of reactants

- $2\,C_2H_6 + 7\,O_2 \longrightarrow 4\,CO_2 + 6\,H_2O$: 60 + 224 = 284 g = 176 + 108 g
- a gas escapes in an open vessel; in a closed one no change
- mass and energy are two separate laws; $E = mc^2$ loss ~1 part in 10^{10} for burning methane

Definite (constant) proportions (Proust): a pure compound has the same elements in the same proportion by mass, whatever its source or method

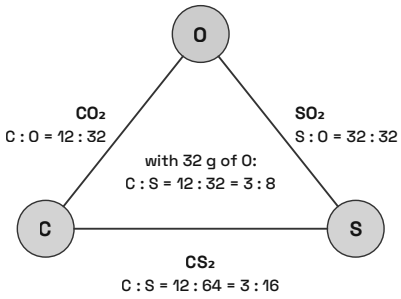
- water H : O = 1 : 8; CO_2 from C or CH_4: C : O = 3 : 8; $p/q = x/y$
- limits: proteins, polymers; non-stoichiometric compounds (berthollides: wüstite $Fe_{0{\cdot}83–0{\cdot}95}O$); isotopes ($D_2O \approx 20\%$ H)

Multiple proportions (Dalton): two elements form several compounds → masses of one with a fixed mass of the other in a small whole-number ratio

- O with 14 g N: N_2O 8, NO 16, NO_2 32, N_2O_5 40 → **1 : 2 : 4 : 5**
- CO (42.9% C) and CO_2 (27.3% C): O per 12 g C = 16 : 32 = 1 : 2

Reciprocal (equivalent) proportions (Richter): A and B each with a fixed mass of C → their mass ratio is the same as, or a simple multiple of, the ratio in which A and B combine

- Cu fixed: S : O = 32 : 16 = 2 : 1; in SO_3 S : O = 2 : 3; (2/1) ÷ (2/3) = 3
- O fixed (CO_2, SO_2): C : S = 12 : 32 = 3 : 8; in CS_2 12 : 64 = 3 : 16 (×2)



C and S each with O (CO_2, SO_2) and with each other (CS_2).

Gay-Lussac's gaseous volumes (1808): reacting and product gases in a small whole-number ratio of volumes at the same T, P

- $2\,H_2(g) + O_2(g) \longrightarrow 2\,H_2O(g)$: 2 : 1 : 2; volumes need not add up (3 → 2)

Avogadro's law

Equal volumes of all gases at the same temperature and pressure contain equal numbers of molecules: $V \propto N$

- 1 vol H_2 + 1 vol Cl_2 → 2 vol HCl ⇒ $H_2 + Cl_2 \longrightarrow 2\,HCl$: elements' molecules are **diatomic**
- **Berzelius** (equal volumes, equal **atoms**) rejected: it would need half atoms

Applications

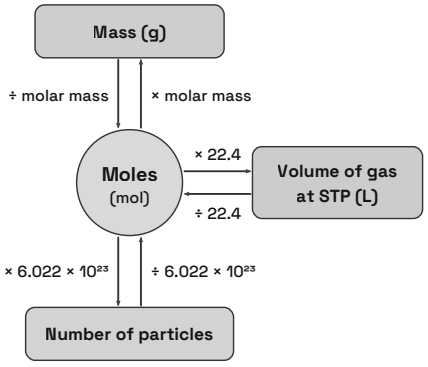
- H_2, O_2, N_2, Cl_2 are diatomic
- $M_r = 2 \times VD$
- molar volume 22.4 L at STP
- 22.4 L holds 6.022×10^{23} molecules
- molecular formula of a gas from its volumes
- explains Gay-Lussac's law

$M_r = 2 \times VD$

1. VD = mass of volume V of gas ÷ mass of V of H_2
2. same V → same number n of molecules: VD = mass of 1 molecule ÷ mass of 1 H_2 molecule
3. H_2 = 2 H atoms: VD = mass of 1 molecule ÷ (2 × mass of 1 H atom) = $M_r/2$

- CO_2: VD 22, heavier than air (≈ 29): gathers in pits and wells
- $M = 2 \times$ (mass of 1 L) ÷ 0.0899 = 22.4 × (mass of 1 L) → molar mass = mass of 22.4 L at STP
- STP here 0 °C and 1 atm (22.4 L); IUPAC's 1 bar gives 22.7 L; use the question's value

The mole



Mass ÷ molar mass, gas volume at STP ÷ 22.4 L, particles ÷ 6.022×10^{23}: all give moles; multiply to go back.

- 12 g C and 63.5 g Cu: 1 mol each, same number of atoms

- 1 mol O_2 = 32 g = 22.4 L at STP = 6.022×10^{23} molecules = 1.204×10^{24} atoms
- 0.5 mol O_2 = 1 mol O atoms; 44.8 L at STP = 2 mol = 1.204×10^{24} molecules
- 1.8 mL water = 1.8 g = 0.1 mol
- equal moles → equal particles; 1 mol O_2 and 1 mol O_3: same volume
- every calculation: given → **moles** → asked for

Calculations from equations

- coefficients give the ratio of **moles** (and gas volumes), never of masses
- $2\,H_2 + O_2 \longrightarrow 2\,H_2O$: moles 2 : 1; masses 4 : 32 = 1 : 8
- $CH_4 + 2O_2$: 0.5 mol CH_4 needs 1 mol O_2
- 4.8 g Mg (0.2 mol) + excess HCl → 0.2 mol H_2 = 4.48 L at STP
- 28 g N_2 (1 mol) with excess H_2 → 2 mol NH_3 = 34 g; 30 g obtained → 88.2%

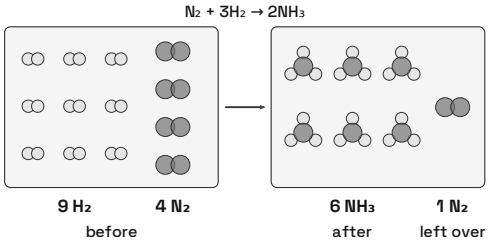
Limiting reactant

1. change each reactant to moles
2. divide each by its coefficient: the **smaller** is limiting
3. product from the limiting reactant; excess left = given − used

- $N_2 + 3\,H_2 \longrightarrow 2\,NH_3$, 9 mol H_2 + 4 mol N_2: H_2 9/3 = 3, N_2 4/1 = 4 → **H_2 limiting**: 6 mol NH_3 (102 g), 1 mol N_2 (28 g) left
- 4 g H_2 (2 mol) + 40 g O_2 (1.25 mol) → H_2 limiting

Why it matters

- the most product a reaction can give
- no waste of chemicals
- fewer unwanted by-products
- cheaper, greener: the cheaper reactant is put in excess



9 mol H_2 and 4 mol N_2: every 3 H_2 take 1 N_2; the H_2 runs out after 3 N_2; 1 N_2 is left.

Yield

- lime kiln: $CaCO_3 \xrightarrow{\Delta} CaO + CO_2$, 100 g → 56 g; 1 kg → 560 g theoretical; 500 g obtained → **89.29%**

Why actual < theoretical

- reaction incomplete (stops partway, as Haber's, or too slow)
- side reactions
- product lost while separating, washing, drying
- impure reactants
- over 100% → product still wet or impure

Formula from percentages

1. % of each element = grams in 100 g
2. ÷ atomic mass → moles
3. ÷ the smallest
4. near .5, .33, .25: × 2, 3, 4 (1 : 1.5 → 2 : 3)
5. empirical formula; molecular = $(EF)_n$

- **from burning**: C = mass of $CO_2 \times 12/44$; H = mass of $H_2O \times 2/18$; O = sample − C − H
- **hydrate**: $Na_2CO_3 \cdot xH_2O$, 2.714 g → 1.006 g Na_2CO_3: water lost ÷ 18 against Na_2CO_3 ÷ 106 gives x

COMMON MISTAKES

✗ coefficients give the ratio of masses
✓ of **moles** (and gas volumes); change grams to moles first.

✗ the reactant with less mass is limiting
✓ compare moles ÷ coefficient.

✗ conservation of "mass and energy"
✓ in chemistry mass is conserved; energy is a separate law.

✗ Berzelius: equal volumes, equal atoms
✓ equal **molecules** (Avogadro); atoms would have to be halved.

✗ 44.8 L at STP holds 1.24×10^{23} molecules
✓ 2 mol = 1.204×10^{24}.

✗ Avogadro constant 6.60×10^{23}
✓ 6.022×10^{23}.

✗ RTP 24.0 L at 25 °C
✓ 24.5 L at 25 °C (24.0 L is at 20 °C); use the question's value.

✗ 500 g is the theoretical yield of the kiln
✓ 560 g theoretical, 500 g experimental.

✗ Haber process at low temperature
✓ about 400–450 °C, iron catalyst, high pressure.