

ESAT

Chemistry

Concentration, gas volumes and yield

Free lesson notes

Covers

C4.8 to C4.11

Format

Lesson, worked examples, self-check

Price

Free

About this lesson

The mole links a mass to a number of particles. This topic extends it to the three other things a chemist actually measures: the *volume of a gas*, the *concentration of a solution*, and *how much product you really got*. Each is a short step on from $n = m/M_r$, and together they let you answer almost any quantitative question in the ESAT.

We cover: the molar gas volume (24 dm^3 per mole at room temperature and pressure), concentration in mol/dm^3 and g/dm^3 , dilution, solubility curves, titration and its calculations, and percentage yield. It assumes the previous topic (moles and reacting masses).

Relative atomic masses used: H = 1, C = 12, N = 14, O = 16, Na = 23, Mg = 24, S = 32, Cl = 35.5, Ca = 40. The real ESAT gives these in the question, and has no calculator, so the numbers are always friendly. Remember the unit conversions: $1 \text{ dm}^3 = 1000 \text{ cm}^3$, so $25.0 \text{ cm}^3 = 0.0250 \text{ dm}^3$.

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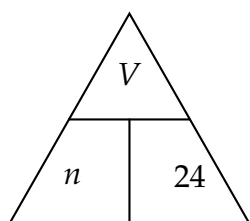
1 The molar gas volume (C4.8)

In a gas the particles are far apart, and their size hardly matters: what fixes the volume is how many particles there are, plus the temperature and the pressure. This is **Avogadro's law**: equal volumes of any gases, at the same temperature and pressure, contain equal numbers of molecules. So one mole of *any* gas takes up the same volume. At room temperature and pressure (rtp, about 20°C and 1 atmosphere) that volume is

$$V_m = 24 \text{ dm}^3/\text{mol} \quad (= 24\,000 \text{ cm}^3/\text{mol}).$$

That is why 2 g of hydrogen and 44 g of carbon dioxide, a mole each, occupy the same space: the heavy molecules are not bigger gaps, just heavier. The relationship is a proportionality, so:

$$n = \frac{V}{24} \quad V = n \times 24 \quad (V \text{ in dm}^3).$$



$$n = \frac{V}{24}$$

$$V = n \times 24$$

V in dm^3 at rtp
($1 \text{ dm}^3 = 1000 \text{ cm}^3$)

If V is given in cm^3 , divide by 24 000 instead. To connect a gas volume with a mass, go through moles: volume $\leftrightarrow$ moles $\leftrightarrow$ mass.

Worked example: volume of a gas

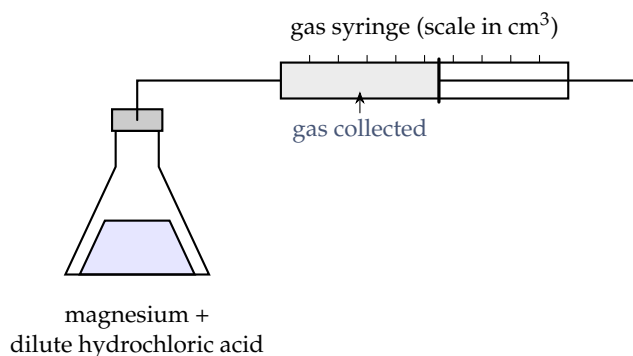
What volume does 8.8 g of carbon dioxide occupy at rtp?

$M_r(\text{CO}_2) = 44$, so $n = 8.8/44 = 0.20 \text{ mol}$. $V = 0.20 \times 24 = 4.8 \text{ dm}^3$.

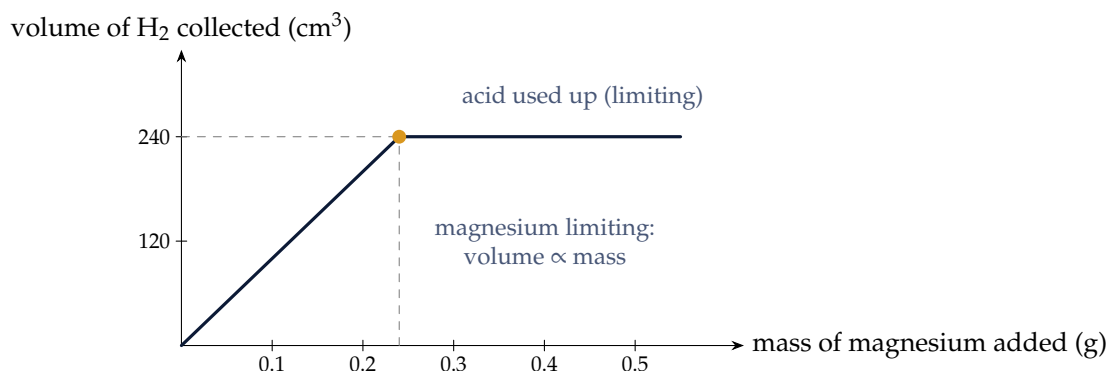
Worked example: gas produced in a reaction

Magnesium reacts with excess dilute acid: $\text{Mg} + 2\text{HCl} \longrightarrow \text{MgCl}_2 + \text{H}_2$. What volume of hydrogen, in cm^3 , is made from 0.24 g of Mg?

$n(\text{Mg}) = 0.24/24 = 0.010 \text{ mol}$. The ratio Mg : H_2 is 1 : 1, so 0.010 mol of H_2 . $V = 0.010 \times 24\,000 = 240 \text{ cm}^3$.



A gas syringe like the one above measures the volume of gas given off, so the reaction itself tells you how many moles were produced. Plotting volume against mass of magnesium (with excess acid) gives a straight line through the origin: more metal, more gas. Once the acid runs out, the line goes flat. The *corner* marks the point where the two reactants are in exactly the right proportion; the plateau is set by the reactant that is limiting.



Common mistakes

- Using 24 when the volume is in cm³: it is 24 000 cm³/mol.
- Using the molar gas volume for solids and liquids. It applies to gases only.
- Using a gas's *atomic* formula: oxygen gas is O₂ (*M_r* = 32), so 16 g is 0.5 mol, i.e. 12 dm³.

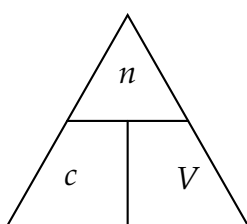
2 Concentration of a solution (C4.9)

A **solution** is a solute dissolved in a solvent (usually water). Its **concentration** says how much solute is in a given volume. Two units are used:

- **mol/dm³**: moles of solute in 1 dm³ of solution. This is the one chemists prefer, because moles are what appear in equations.
- **g/dm³**: mass of solute in 1 dm³ of solution.

The definition is the same as any "amount per volume":

$$c = \frac{n}{V} \quad n = c \times V \quad V = \frac{n}{c} \quad (V \text{ in dm}^3).$$



$$c = \frac{n}{V}$$

$$n = c \times V$$

$$V = \frac{n}{c}$$

c in mol/dm³
V in dm³
 (cm³ ÷ 1000)

Always convert volumes to dm^3 first: $25.0 \text{ cm}^3 \rightarrow 0.0250 \text{ dm}^3$ (divide by 1000). To change between the two units, use M_r : $\text{g/dm}^3 = \text{mol/dm}^3 \times M_r$. For example 0.10 mol/dm^3 of NaOH ($M_r = 40$) is 4.0 g/dm^3 .

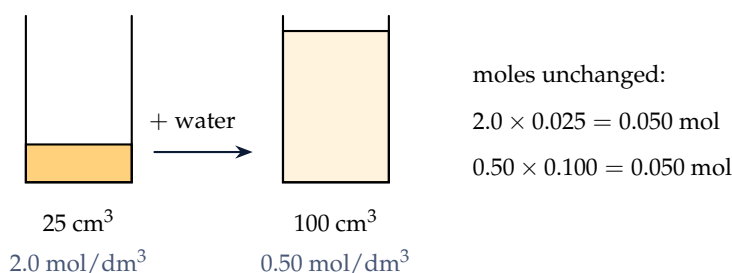
Worked example: concentration from a mass

4.0 g of sodium hydroxide is dissolved in water to make 500 cm^3 of solution. Find the concentration in mol/dm^3 .
 $M_r(\text{NaOH}) = 40$, so $n = 4.0/40 = 0.10 \text{ mol}$. $V = 0.500 \text{ dm}^3$. $c = 0.10/0.500 = 0.20 \text{ mol/dm}^3$.

Worked example: moles in a sample

How many moles of acid are in 25.0 cm^3 of a 0.20 mol/dm^3 solution?
 $n = c \times V = 0.20 \times 0.0250 = 0.0050 \text{ mol}$.

Dilution. Adding water to a solution does not change the number of moles of solute, only the volume. So n before = n after: $c_1 V_1 = c_2 V_2$.



Worked example: dilution

25 cm^3 of 2.0 mol/dm^3 acid is made up to 100 cm^3 with water. What is the new concentration?
 $n = 2.0 \times 0.025 = 0.050 \text{ mol}$, unchanged. $c = 0.050/0.100 = 0.50 \text{ mol/dm}^3$. (The volume is four times larger, so the concentration is four times smaller.)

Common mistakes

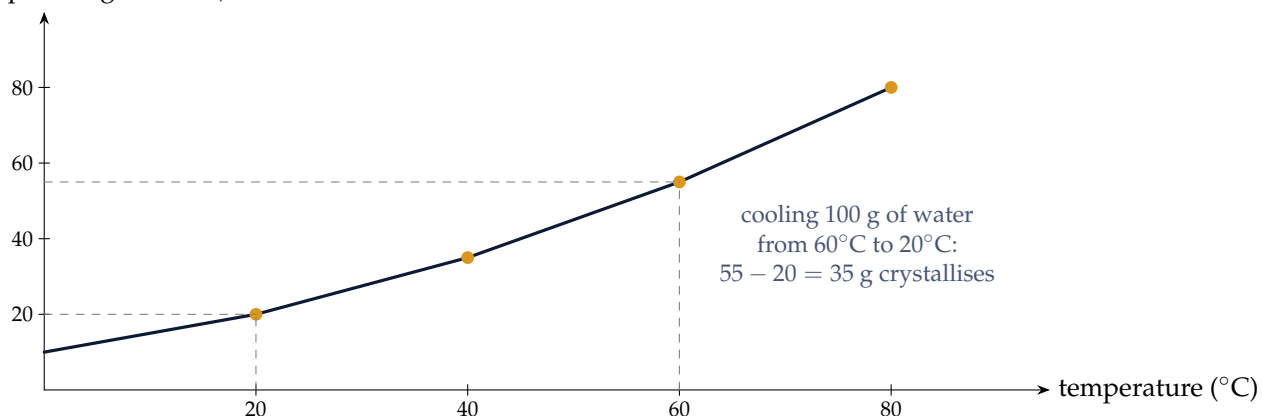
- Forgetting to convert cm^3 to dm^3 : this gives answers 1000 times too big.
- Thinking the *volume of water added* is the final volume: in the example above, 75 cm^3 of water is added to make 100 cm^3 .
- Mixing the two units: a concentration of 4.0 g/dm^3 is *not* 4.0 mol/dm^3 .

3 Solubility (C4.9)

The **solubility** of a substance is the maximum mass of it that dissolves in 100 g of water at a given temperature. A solution holding that maximum is **saturated**: any more solute stays undissolved. For most solids, solubility *increases* with temperature, so a hot saturated solution that cools can hold less and the extra comes out as **crystals**.

A **solubility curve** plots solubility against temperature. Read values straight from the graph, using the guide lines, and scale by the mass of water: if the solubility is 55 g per 100 g of water, then 200 g of water dissolves 110 g .

solubility (g per 100 g of water)



Worked example: crystals on cooling

Use the graph above. A saturated solution is made with 200 g of water at 60 °C. How much solid crystallises when it is cooled to 20 °C?

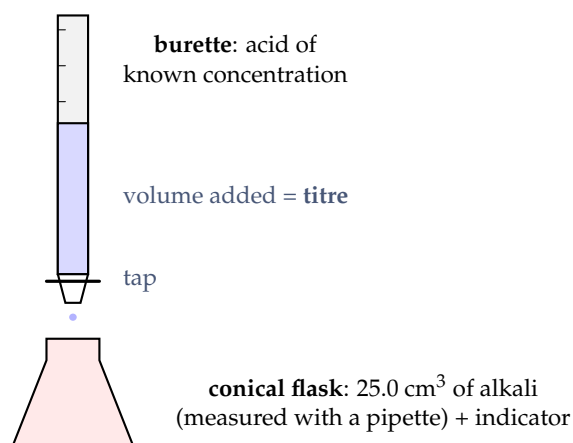
At 60 °C, 200 g of water holds $2 \times 55 = 110$ g. At 20 °C it holds only $2 \times 20 = 40$ g. The difference crystallises: $110 - 40 = 70$ g.

Common mistakes

- Using the mass of the whole solution instead of the mass of *water*: solubility is per 100 g of water.
- Subtracting the wrong way round, or forgetting to scale up for the amount of water.
- Assuming that all solids are more soluble when hot. Gases, such as oxygen, become *less* soluble as the temperature rises.

4 Titration: the method (C4.10)

How do you find the concentration of an unknown solution? You react it, exactly, with a solution of *known* concentration, and measure the volumes. This is a **titration**, most often used for an acid and an alkali.



1. Use a **pipette** to measure exactly 25.0 cm^3 of one solution into a conical flask. (A pipette is more accurate than a measuring cylinder.)
2. Add a few drops of an **indicator** (for example methyl orange or phenolphthalein), which changes colour at the end-point.
3. Fill the **burette** with the other solution and record the starting level. Add it to the flask slowly, swirling, until the indicator *just* changes colour: the **end-point**, where the reactants are in exactly the ratio of the equation.
4. The **titre** is the volume added from the burette (final reading minus start reading).

The first run is a **rough** titration to find roughly where the end-point is. Then repeat carefully until you get two or more

concordant results (agreeing within 0.1 cm^3), and take the mean of the concordant titres only. Never include the rough titre or a result that is far from the others.

rough	run 1	run 2	run 3
21.0	20.2	20.0	20.0

concordant titres (within 0.1 cm^3) shaded:
 $\text{mean} = \frac{20.0 + 20.0}{2} = 20.0 \text{ cm}^3$

Worked example: choosing the titre

Titres: rough 21.0, then 20.2, 20.0 and 20.0 cm^3 . What titre should be used?
 The rough titre is ignored. 20.0 and 20.0 cm^3 agree within 0.1 cm^3 and are concordant; 20.2 is not. $\text{Mean} = (20.0 + 20.0)/2 = 20.0 \text{ cm}^3$.

Common mistakes

- Averaging all the readings, including the rough one. Only concordant titres are averaged.
- Reading the burette from the top: a burette scale increases *downwards*, and you subtract start from finish.
- Using a measuring cylinder for the fixed volume: a pipette is the accurate instrument.

5 Titration calculations (C4.10)

A titration calculation is the mole method in a fixed order. This is the routine to follow every time.

1. Write the balanced equation.
2. Find the moles of the solution you know completely: $n = c \times V$ (with V in dm^3).
3. Use the mole ratio in the equation to find the moles of the other reactant.
4. Find the unknown quantity: $c = n/V$ for the unknown solution (using its volume in dm^3).

Worked example: acid and alkali (1 : 1)

25.0 cm^3 of sodium hydroxide solution is neutralised by 20.0 cm^3 of 0.10 mol/dm^3 hydrochloric acid. Find the concentration of the alkali. $\text{NaOH} + \text{HCl} \longrightarrow \text{NaCl} + \text{H}_2\text{O}$
 $n(\text{HCl}) = 0.10 \times 0.0200 = 0.0020 \text{ mol}$. The ratio is 1 : 1, so $n(\text{NaOH}) = 0.0020 \text{ mol}$. $c = 0.0020/0.0250 = 0.080 \text{ mol/dm}^3$.

Worked example: acid and alkali (1 : 2)

25.0 cm^3 of 0.10 mol/dm^3 sodium hydroxide needs 20.0 cm^3 of sulfuric acid to neutralise it: $\text{H}_2\text{SO}_4 + 2 \text{NaOH} \longrightarrow \text{Na}_2\text{SO}_4 + 2 \text{H}_2\text{O}$. Find the acid concentration.
 $n(\text{NaOH}) = 0.10 \times 0.0250 = 0.0025 \text{ mol}$. The equation says 1 H_2SO_4 reacts with 2 NaOH , so $n(\text{H}_2\text{SO}_4) = 0.0025/2 = 0.00125 \text{ mol}$. $c = 0.00125/0.0200 = 0.0625 \text{ mol/dm}^3$.

Using a titre to find a mass or purity. If the unknown sample was dissolved and made up to a larger volume, take care: the pipette removes only part of it. Scale back up by the ratio of the volumes. In an impure sample, the moles found tell you the mass of the pure substance; divide by the sample mass for the percentage purity.

Common mistakes

- Ignoring the mole ratio: for H_2SO_4 with NaOH the ratio is 1 : 2, and using 1 : 1 gives an answer a factor of 2 out.
- Using the volume of the wrong solution in $c = n/V$. The concentration you want belongs to the solution whose volume you divide by.
- Using cm^3 for one volume and dm^3 for another.

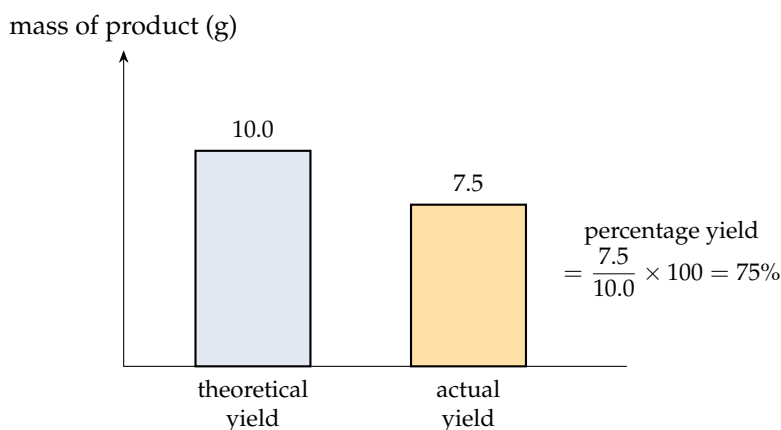
Exam tip

If both volumes are in cm^3 , you can leave them there for a 1 : 1 reaction, because $c_1 V_1 = c_2 V_2$ and the factor of 1000 cancels. For other ratios the conversion is still needed to get moles.

6 Percentage yield (C4.11)

A balanced equation tells you the **theoretical yield**: the mass of product you would get if every reactant turned into product. In reality you get less: the reaction may be reversible, some product is lost in transfers and filtering, there are side reactions, or the reactants are impure. The mass you really obtain is the **actual yield**.

$$\text{percentage yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100.$$



Method. (1) Work out the theoretical yield from the *limiting* reactant, using moles, as in the last topic. (2) Divide the actual yield by it and multiply by 100. The theoretical yield is the maximum; a percentage yield above 100% is impossible and means an error (often a wet product, or a wrong calculation).

When there are several steps in a process, the yields *multiply*: two steps of 80% and 50% give an overall yield of $0.80 \times 0.50 = 40\%$.

Worked example: percentage yield

10.0 g of calcium carbonate is heated: $\text{CaCO}_3 \rightarrow \text{CaO} + \text{CO}_2$. 4.2 g of calcium oxide is collected. Find the percentage yield.

$n(\text{CaCO}_3) = 10.0/100 = 0.100$ mol, giving 0.100 mol of CaO. Theoretical mass = $0.100 \times 56 = 5.6$ g. Yield = $\frac{4.2}{5.6} \times 100 = 75\%$.

Worked example: working back from a yield

A reaction has a yield of 60%. The theoretical yield is 20 g. What mass is actually made?

Actual = $\frac{60}{100} \times 20 = 12$ g.

Common mistakes

- Using the mass of the reactant (instead of the *theoretical product*) as the denominator.
- Calculating the theoretical yield from the reactant that is in excess instead of the limiting reactant.
- Inverting the fraction: the actual yield is the smaller number, so the answer should be less than 100%.

7 Putting it together

Harder questions chain these ideas. Almost all of them follow one skeleton:

mass or volume or concentration $\rightarrow$ moles $\rightarrow$ (equation ratio) $\rightarrow$ moles $\rightarrow$ the thing asked for

The conversions into and out of moles are m/M_r , $V/24$ (gas) and $c \times V$ (solution). A **limiting reactant** adds a step: compare moles divided by coefficients. A **percentage yield** adds one at the end. A **dilution** or a **sample taken from a bigger volume** adds a scaling factor.

An example with three conversions. 1.2 g of magnesium is added to 100 cm³ of 0.50 mol/dm³ hydrochloric acid. What volume of hydrogen forms at rtp? Moles: Mg $1.2/24 = 0.050$; HCl $0.50 \times 0.100 = 0.050$. The equation $\text{Mg} + 2\text{HCl} \rightarrow$

$\text{MgCl}_2 + \text{H}_2$ needs two HCl per Mg, so the acid is limiting: $0.050/2 = 0.025$ mol of H_2 . $V = 0.025 \times 24 = 0.60 \text{ dm}^3$ (600 cm^3).

Worked example: limiting reactant and gas volume

2.4 g of magnesium is added to 200 cm^3 of 0.50 mol/dm^3 hydrochloric acid. Calculate the volume of hydrogen at rtp.

Mg: $2.4/24 = 0.10$ mol. HCl: $0.50 \times 0.200 = 0.10$ mol. HCl needs twice as many moles as Mg, so 0.10 mol of HCl is only enough for 0.05 mol of Mg: HCl limits. $n(\text{H}_2) = 0.10/2 = 0.050$ mol. $V = 0.050 \times 24 = 1.2 \text{ dm}^3$.

Common mistakes

- Taking the first reactant given as the limiting one. Always compare moles divided by coefficients.
- Dropping a step: write the chain of conversions on one line before calculating.
- Changing the units half-way: choose dm^3 for both gas and solution volumes and stay there.

Check yourself

Try these without looking back, then compare with the answers on the next page.

1. What volume does 0.5 mol of gas occupy at rtp?
2. What volume of H_2 is made from 0.12 g of Mg and excess acid?
3. Find the concentration when 0.2 mol is dissolved to make 500 cm^3 .
4. 50 cm^3 of 1.0 mol/dm^3 acid is made up to 250 cm^3 . New concentration?
5. What is the concentration in g/dm^3 of 0.10 mol/dm^3 NaOH?
6. Titres: 22.8, 21.9, 21.4, 21.4. Which value is used?
7. A reaction should give 40 g but gives 30 g. Percentage yield?
8. Two steps have yields of 90% and 50%. Overall yield?

Answers to the self-check

1. $0.5 \times 24 = 12 \text{ dm}^3$.
2. $0.12/24 = 0.005 \text{ mol}$, $\times 24\,000 = 120 \text{ cm}^3$.
3. $0.2/0.5 = 0.4 \text{ mol/dm}^3$.
4. $1.0 \times 50/250 = 0.20 \text{ mol/dm}^3$.
5. $0.10 \times 40 = 4.0 \text{ g/dm}^3$.
6. 21.4 (the concordant pair).
7. $30/40 = 75\%$.
8. $0.90 \times 0.50 = 45\%$.

Next step

Heurisco's *ESAT Chemistry: Concentration, Gas Volumes and Yield Practice Set* has twenty original exam-style questions on this lesson, with graphs and data tables, in order of difficulty, with a full worked solution and a note on the trap in each question.