Topic 3: Quantitative chemistry

AQA GCSE Chemistry (8462), topic 4.3

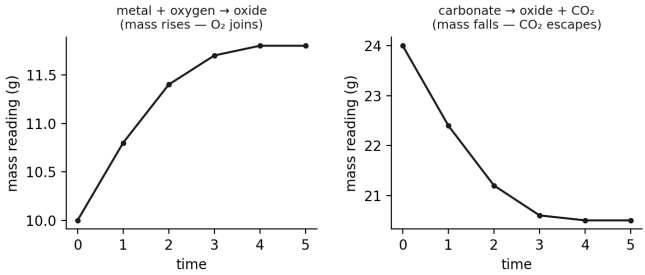
Quantitative chemistry: counting atoms by mass

- use **conservation of mass** to explain why some reactions seem to gain or lose mass;
- calculate relative formula masses and percentages by mass;
- (higher tier) convert between mass, moles and numbers of particles, and use moles to find reacting masses and limiting reactants;

Quantitative chemistry: counting atoms by mass (2)

- calculate concentrations in g/dm^3 and (higher tier) mol/dm^3;
- (chemistry only) calculate **percentage yield**, **atom economy** and gas volumes.

Figure: Quantitative chemistry: counting atoms by mass



- Metal gains oxygen from air.
- **Carbonate** releases CO_2 to air.
- Account for all reactants and products.

Conservation of mass and equations

- The law of **conservation of mass** says that atoms are not made or destroyed in a chemical reaction.
- They are only rearranged, so the total mass of the products equals the total mass of the reactants.
- A symbol equation must therefore be **balanced** : the same number of each kind of atom on both sides.
- There are two kinds of number in an equation.

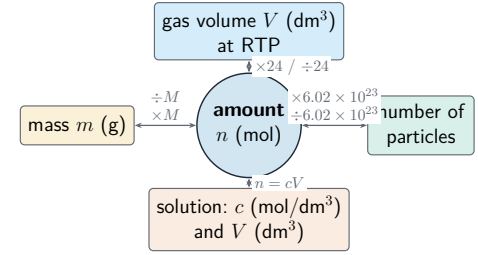
$$\text{percentage by mass} = \frac{A_r \times \text{number of atoms of the element}}{M_r} \times 100$$

Moles and reacting masses (higher tier)

- Atoms are far too small to count one by one, so chemists count them in moles.
- One **mole** of a substance contains 6.02×10^{23} particles: this number is the **Avogadro constant** .
- The mass of one mole in grams is equal to the Mr, so one mole of carbon atoms (Ar 12) weighs 12 g and one mole of water (Mr 18) weighs 18 g.
- This mass per mole is the **molar mass** , in g/mol.

$$\text{moles} = \frac{\text{mass (g)}}{M_r}$$

Figure: Moles and reacting masses (higher tier)



the equation ratio links amounts in moles, never masses

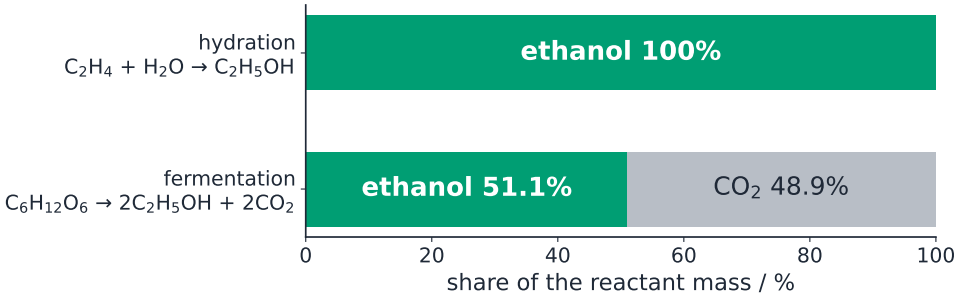
- Mass ÷ M gives moles; moles × M gives mass.
- Use the equation's ratio in moles.

Yield and atom economy (chemistry only)

- **Percentage yield.** You almost never get the full amount of product that the equation predicts, the **theoretical mass** .
- The reaction may be **reversible** , so it does not go to completion; some product is lost when it is separated; and some reactants may react in other ways.
- **Worked example 9.** The **theoretical mass** is 8.0 g and you collect 6.0 g of dry product.
- The **percentage yield** is $6.0/8.0 \times 100 = 75\%$.

$$\text{percentage yield} = \frac{\text{mass of product actually made}}{\text{maximum theoretical mass}} \times 100$$

Figure: Yield and atom economy (chemistry only)



- **Atom economy** = mass of desired product ÷ total reactant mass × 100.
- Addition reactions waste nothing.

Concentrations in mol/dm^3 (chemistry only, higher tier)

- **Molar concentration** is the number of moles of **solute** in each dm^3 of solution:
- To change it into **mass concentration** , multiply by the Mr: 0.20 mol/dm^3 NaOH is $0.20 \times 40 = 8.0$ g/dm^3.
- In a **titration** , you measure the exact volumes of an acid and an alkali that react.
- If you know one **concentration**, you can calculate the other, using the mole ratio in the equation.

$$\text{moles} = \text{concentration (mol/dm}^3\text{)} \times \text{volume (dm}^3\text{)}$$

Gas volumes (chemistry only, higher tier)

- Equal volumes of different gases at the same temperature and pressure contain the same number of molecules.
- At **room temperature and pressure** (20 °C and 1 atmosphere), one mole of any gas takes up 24 dm^3.
- **Worked example 12.** What volume does 4.4 g of carbon dioxide take up at **room temperature and pressure**?
- Moles $= 4.4/44 = 0.100$ mol, so volume $= 0.100 \times 24 = 2.40$ dm^3, or 2400 cm^3.

$$\text{volume of gas } (dm^3) = \text{moles} \times 24$$

Sources for this lesson

- AQA, *GCSE Chemistry (8462) specification*, section 4.3.
- G. H. Lockwood, *Final Report of the Board of Inquiry Investigating the Circumstances of an Accident Involving the Air Canada Boeing 767 Aircraft C-GAUN that Effected an Emergency Landing at Gimli, Manitoba* (Government of Canada, April 1985).
- A. Lavoisier, *Traité élémentaire de chimie* (1789), where the law of **conservation of mass** was first stated clearly.
- International Bureau of Weights and Measures, *The International System of Units* (9th edition, 2019), on the mole and the **Avogadro constant**.

Common errors

- Changing a **subscript** to balance an equation; change only the numbers in front of formulae.
- Forgetting to multiply by the number in front of a formula when adding up masses.
- Saying mass is "lost" when a **carbonate** is heated; the carbon dioxide escapes into the air.

Common errors (2)

- Using 16 instead of 32 for oxygen gas, which is made of O_2 molecules.
- Using cm^3 in a formula that needs dm^3; divide by 1000 first.
- Calculating the product from the reactant in **excess** instead of the **limiting reactant**.
- Assuming the moles of acid equal the moles of alkali in every **titration**; use the ratio in the equation.

Key terms

- **Conservation of mass**: in a chemical reaction, no atoms are made or destroyed, so total mass does not change.
- **Relative formula mass (Mr)**: the sum of the relative atomic masses of the atoms in a formula.
- **Mole**: the amount of substance that contains 6.02×10^{23} particles.

Key terms (2)

- **Limiting reactant**: the reactant that is used up first and so limits the amount of product.
- **Concentration**: the amount of **solute** in each dm^3 of solution.
- **Percentage yield**: the mass of product made as a percentage of the maximum **theoretical mass**.
- **Atom economy**: the mass of the desired product as a percentage of the total mass of the reactants.

Check yourself

- 1.
- Calculate the Mr of calcium hydroxide, $Ca(OH)_2$ (Ca = 40, O = 16, H = 1). 2.
- Why does the mass of an open flask fall when marble chips react with acid? 3. (HT) How many moles are in 9.0 g of water? 4. (HT) 2.4 g of magnesium reacts with **excess** hydrochloric acid: $Mg + 2HCl \rightarrow MgCl_2 + H_2$.
- What mass of hydrogen is made? 5. 5.0 g of salt is dissolved to make 200 cm^3 of solution.