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Topic 4.5: Stoichiometry

Revision Notes · AP Chemistry

UNIT 4 · CHEMICAL REACTIONS

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Coefficients Are Mole Ratios, Not Mass Ratios

Stoichiometry is the arithmetic of a balanced equation: it converts an amount of one substance in a reaction into the amount of any other.

A balanced equation is a recipe written in moles. Its coefficients count particles, and so moles, never grams. In $N_2 + 3H_2 \rightarrow 2NH_3$, the ratio 1 : 3 : 2 says that 1 mol of N_2 reacts with 3 mol of H_2 to make 2 mol of NH_3.

The ratio works in either direction and between any pair of substances. From 6.0 mol of H_2 with excess N_2, the H_2 to NH_3 ratio gives $6.0 \times 2/3 = 4.0$ mol of NH_3.

Only coefficients give these ratios. Subscripts belong to a formula and fix what the substance is: changing H_2O into H_2O_2 would change the substance, not the amount.

Every conversion factor taken from an equation is built the same way: coefficient of the substance wanted, divided by coefficient of the substance given.

The Calculation Route: Mass, Moles, Ratio, Moles, Mass

The equation speaks only in moles, so a mass must be translated before the ratio can be used, and the answer translated back afterwards.

Step one: convert the given mass into moles by dividing by the molar mass. Step two: apply the mole ratio from the balanced equation. Step three: convert the resulting moles into a mass by multiplying by the molar mass of the substance wanted.

Molar masses are built from the periodic table: $CaCO_3$ is $40.08 + 12.01 + 3 \times 16.00 = 100.09$ g/mol, and CO_2 is 44.01 g/mol.

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Keep full unrounded values through all three steps and round once at the end, normally to three significant figures. Rounding at every step lets small errors grow.

When Reactants Run Out: Limiting and Excess

Real mixtures seldom match the recipe exactly. The reaction stops the moment one reactant is used up, and that reactant is the limiting reactant: it alone sets how much product can form.

The reliable test is to calculate, from each reactant separately, how much product it could make. The reactant that makes less product is the limiting one; every other reactant is in excess, and some of it is left over.

Comparing the starting masses, or even the starting moles, without the coefficients misleads. In $2Mg + O_2 \rightarrow 2MgO$, oxygen can be present in fewer moles than magnesium yet still be in excess, because each mole of O_2 consumes two moles of Mg.

The amount of an excess reactant left over is found by subtraction: what was there at the start, minus what the limiting reactant consumed, worked out through the mole ratio.

Theoretical Yield, Actual Yield and Percent Yield

The theoretical yield is the maximum mass of product the stoichiometry allows, calculated from the limiting reactant with a perfect reaction and no losses. It is a prediction, never a measurement.

The actual yield is what is really obtained and isolated in the laboratory. It falls below the theoretical yield because product is lost on glassware and filter paper, some reactant reacts in side reactions, and some reactions never quite finish.

Percent yield compares the two: percent yield = (actual yield ÷ theoretical yield) × 100. Rearranged, actual yield = percent yield × theoretical yield, so a 75% yield on a theoretical 20.0 g means 15.0 g was isolated.

A yield reported above 100% cannot describe pure, dry product. It signals product that is wet or impure, or a theoretical yield calculated from the wrong reactant.

Stoichiometry in Solution

For a reactant in solution, moles come from concentration and volume: moles = concentration (mol/L) × volume (L). The volume must be in litres, so millilitres are divided by 1000 first.

Solution stoichiometry then follows the same route: moles of one reactant, the equation's ratio, then moles or concentration of the other.

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A neutralisation from the quiz shows the pattern. 25.0 mL of 0.100 M NaOH holds 0.0250 × 0.100 = 0.00250 mol of OH^-. HCl reacts with it in a 1 : 1 ratio, so 0.00250 mol of HCl was present; in 20.0 mL that is 0.00250 ÷ 0.0200 = 0.125 M.

Not every ratio is 1 : 1: one mole of H_2SO_4 neutralises two moles of NaOH, because the equation $H_2SO_4 + 2NaOH \rightarrow Na_2SO_4 + 2H_2O$ says so. The equation, not habit, sets the ratio.

Worked Examples

Example 1: A mass to mass conversion. What mass of CO_2 forms when 25.0 g of $CaCO_3$ decomposes, $CaCO_3 \rightarrow CaO + CO_2$? Moles of $CaCO_3$ = 25.0 ÷ 100.09 = 0.250 mol. The ratio is 1 : 1, so 0.250 mol of CO_2 forms, with mass 0.250 × 44.01 = 11.0 g. The other product checks the arithmetic: 0.250 mol of CaO has mass 0.250 × 56.08 = 14.0 g, and 14.0 + 11.0 = 25.0 g, so mass is conserved.

Example 2: Finding the limiting reactant. 6.00 g of Mg and 4.00 g of O_2 react: $2Mg + O_2 \rightarrow 2MgO$. Moles present: Mg = 6.00 ÷ 24.31 = 0.247 mol; O_2 = 4.00 ÷ 32.00 = 0.125 mol. Product from the Mg: 0.247 mol of MgO. Product from the O_2: 0.125 × 2 = 0.250 mol of MgO. Magnesium makes less, so magnesium is limiting and the yield is 0.247 × 40.31 = 9.95 g of MgO. The O_2 consumed is 0.247 ÷ 2 = 0.123 mol, leaving 0.00159 mol, about 0.0510 g, unreacted.

Example 3: A percent yield. 12.25 g of $KClO_3$ is decomposed, $2KClO_3 \rightarrow 2KCl + 3O_2$, and 4.32 g of O_2 is collected. Moles of $KClO_3$ = 12.25 ÷ 122.55 = 0.100 mol. Oxygen formed = 0.100 × 3/2 = 0.150 mol, a theoretical yield of 0.150 × 32.00 = 4.80 g. Percent yield = 4.32 ÷ 4.80 × 100 = 90.0%.

Traps That Cost Marks in Topic 4.5

Using masses as if they were mole ratios. Coefficients compare moles, not grams: 6.0 g of H_2 does not make 6.0 g of NH_3. Always pass through moles.

Picking the limiting reactant by the smallest mass or the fewest moles. Neither comparison uses the coefficients. With 8.00 g of Mg (0.329 mol) and 6.00 g of O_2 (0.188 mol), the oxygen has both the smaller mass and fewer moles, yet magnesium limits, because the O_2 could make 0.375 mol of MgO against magnesium's 0.329 mol. Only the product test decides.

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Assuming every ratio is 1 : 1. Some are, many are not: 2.0 mol of $KClO_3$ makes 3.0 mol of O_2, not 2.0 mol. Read the coefficients every time.

Calculating the theoretical yield from the excess reactant. That overstates the maximum and drags the percent yield down with it. The limiting reactant, and only the limiting reactant, sets the theoretical yield.

Inverting the percent yield. It is actual divided by theoretical. Turning it over can produce a yield above 100%, which should be an instant warning.

Feeding millilitres into moles = concentration × volume. The formula needs litres. 25.0 mL is 0.0250 L; skipping the conversion inflates the moles a thousandfold.

Key Relationships to Remember

- Mole ratios come only from the coefficients of the balanced equation, never from subscripts and never from masses
- Every stoichiometry calculation follows one route: mass to moles, mole ratio, moles to the wanted amount
- The limiting reactant is the one that produces less product when each reactant is tested separately; it alone sets the theoretical yield
- Excess reactant left over equals the amount started with minus the amount the limiting reactant consumed
- Percent yield = (actual yield ÷ theoretical yield) × 100, and a value above 100% means the product is impure or the calculation is wrong
- For solutions, moles = concentration × volume in litres, and neutralisation ratios still come from the balanced equation

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Topic 4.5 Quiz: Stoichiometry

Topic 4.5 Worksheet

Unit 4: All Topics

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