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Topic 4.5: Stoichiometry

Worksheet · AP Chemistry

UNIT 4 · CHEMICAL REACTIONS

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How to Use This Worksheet

Work in order. Section A checks the ideas: where mole ratios come from, the key definitions, and the basic conversions. Section B builds the skill: mass to mass calculations, limiting reactants, percent yield and solution stoichiometry. Section C combines everything in exam style, including testing claims that sound plausible but fail.

Write the three steps of every calculation out in full: convert to moles, apply the mole ratio from the balanced equation, then convert to what is asked. Carry unrounded values through and round only the final answer, to three significant figures. Molar masses you need: H_2O 18.02, NH_3 17.03, CO_2 44.01, NaCl 58.44, $CaCO_3$ 100.09, $KClO_3$ 122.55, MgO 40.31, $ZnCl_2$ 136.28 g/mol; atomic masses: H 1.008, C 12.01, O 16.00, Mg 24.31, Zn 65.38 g/mol.

Two habits separate correct answers from near misses in this topic. First, never compare reactants by mass or by moles alone when deciding which one limits: test how much product each could make. Second, treat a percent yield above 100% as an alarm bell, never as a success. Finally, whenever both products' masses are known, check them against the reactant masses consumed: conservation of mass catches arithmetic slips.

Do not look at the answer key until you have attempted every question.

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Section A: Recall and Concepts

A1. (a) State where the mole ratio between two substances in a reaction comes from. (b) For $2H_2 + O_2 \rightarrow 2H_2O$, state how many moles of H_2O form from 3.0 mol of O_2, and how many moles of O_2 are needed for 5.0 mol of H_2. (c) For $N_2 + 3H_2 \rightarrow 2NH_3$, calculate the moles of NH_3 formed from 9.0 mol of N_2 with H_2 in excess. (d) Explain why the subscripts in a formula cannot be used as a mole ratio.

A2. Define each term in one sentence. (a) Mole ratio. (b) Limiting reactant. (c) Excess reactant. (d) Theoretical yield. (e) Actual yield. (f) Percent yield, including the formula.

A3. Potassium chlorate decomposes: $2KClO_3 \rightarrow 2KCl + 3O_2$. (a) How many moles of O_2 form from 5.0 mol of $KClO_3$? (b) How many moles of $KClO_3$ are needed to produce 6.0 mol of O_2? (c) How many moles of KCl form alongside the O_2 in part (a)?

A4. (a) Calculate the moles in 11.0 g of CO_2. (b) Calculate the mass of 0.750 mol of NaCl. (c) Calculate the moles in 90.0 g of H_2O. (d) State the two conversion factors that connect mass and moles, and what each one does.

A5. Magnesium burns: $2Mg + O_2 \rightarrow 2MgO$. (a) A mixture holds 4.0 mol of Mg and 4.0 mol of O_2. Which reactant is limiting, and how much O_2 is left over? (b) A second mixture holds 4.0 mol of Mg and 1.0 mol of O_2. Which reactant is limiting now, how much MgO forms, and how much Mg is left over? (c) Explain why "the reactant present in the smaller amount" is not a safe way to choose the limiting reactant.

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Section B: Core Practice

B1. Methane burns completely: $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$. A sample of 10.0 g of CH_4 is burned in excess oxygen. (a) Calculate the moles of CH_4 burned. (b) Calculate the mass of CO_2 produced. (c) Calculate the mass of O_2 consumed.

B2. A ribbon of 8.00 g of Mg is burned in 6.00 g of O_2: $2Mg + O_2 \rightarrow 2MgO$. (a) Calculate the moles of each reactant present. (b) By calculating the MgO each reactant could produce, identify the limiting reactant. (c) Calculate the theoretical yield of MgO. (d) Calculate the mass of the excess reactant left over.

B3. (a) A reaction has a theoretical yield of 25.5 g and an actual yield of 20.4 g. Calculate the percent yield. (b) A reaction has a percent yield of 85.0% and a theoretical yield of 40.0 g. Calculate the actual yield. (c) A student reports a percent yield of 105%. Give two distinct reasons why the isolated product cannot be pure and dry. (d) In part (a), how much more product would need to be recovered to bring the yield up to 90.0%?

B4. Hydrochloric acid is neutralised by sodium hydroxide: $HCl + NaOH \rightarrow NaCl + H_2O$. (a) Calculate the moles of HCl in 25.0 mL of 0.200 M HCl. (b) Calculate the volume of 0.150 M NaOH needed to neutralise exactly the acid in part (a). (c) Sulfuric acid reacts differently: $H_2SO_4 + 2NaOH \rightarrow Na_2SO_4 + 2H_2O$. Calculate the moles of NaOH needed to neutralise 20.0 mL of 0.500 M H_2SO_4. (d) Calculate the volume of 0.400 M NaOH that supplies the moles in part (c). (e) Calculate the mass of NaCl formed in the neutralisation in parts (a) and (b).

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B5. Ammonia is made by $N_2 + 3H_2 \rightarrow 2NH_3$. A reactor is charged with 14.0 g of N_2 and 4.00 g of H_2. (a) Calculate the moles of each reactant. (b) Identify the limiting reactant, showing the test you used. (c) Calculate the theoretical yield of NH_3. (d) The reactor actually produces 12.8 g of NH_3. Calculate the percent yield. (e) Calculate the mass of H_2 left over in the reactor when the limiting reactant is exhausted.

B6. Zinc reacts with excess hydrochloric acid: $Zn + 2HCl \rightarrow ZnCl_2 + H_2$. A 13.1 g piece of zinc is used. (a) Calculate the moles of zinc. (b) Calculate the moles of HCl consumed. (c) Calculate the mass of H_2 produced. (d) Calculate the mass of $ZnCl_2$ produced, and check your answers using conservation of mass. (Molar mass of HCl is 36.46 g/mol.)

Section C: Exam Style

C1. A student adds 20.0 g of $CaCO_3$ to 100 mL of 3.00 M HCl: $CaCO_3 + 2HCl \rightarrow CaCl_2 + H_2O + CO_2$. (a) Calculate the moles of each reactant present. (b) Identify the limiting reactant, showing your working. (c) Calculate the theoretical yield of CO_2. (d) The student collects 5.94 g of CO_2. Calculate the percent yield. (e) Calculate the mass of $CaCO_3$ left unreacted.

C2. Propane burns: $C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$. A burner is supplied with 11.0 g of C_3H_8 and 30.0 g of O_2. (a) Calculate the moles of each reactant. (b) Identify the limiting reactant, showing your working. (c) Calculate the masses of CO_2 and of H_2O produced. (d) Calculate the mass of the excess reactant left over.

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C3. Three students make claims about the reaction $2Mg + O_2 \rightarrow 2MgO$. Evaluate each claim. Claim 1: "With 8.00 g of Mg and 6.00 g of O_2, the oxygen must be the limiting reactant, because it has the smaller mass." Claim 2: "My percent yield was 110%, so my reaction worked better than the theory predicted." Claim 3: "If I mix exactly 2.0 mol of Mg with 2.0 mol of O_2, neither reactant is in excess, because the amounts are equal." (a) Evaluate Claim 1 by carrying out the limiting reactant test. (b) Evaluate Claim 2, explaining what a yield above 100% really indicates. (c) Evaluate Claim 3 by working out what actually happens to each reactant.

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Answer Key

A1. (a) From the coefficients of the balanced equation: the ratio of the coefficients of the two substances. (b) 6.0 mol of H_2O (3.0 × 2/1), and 2.5 mol of O_2 (5.0 × 1/2). (c) 9.0 × 2/1 = 18.0 mol of NH_3. (d) Subscripts state the fixed composition of one formula unit, for example the two H atoms inside each H_2O molecule. They describe what a substance is, not how much of it reacts; only the coefficients in front of the formulas give reacting amounts.

A2. (a) The ratio between the amounts in moles of any two substances in a reaction, taken from the coefficients of the balanced equation. (b) The reactant that is completely consumed first, stopping the reaction and setting the maximum amount of product. (c) A reactant supplied in more than the amount the limiting reactant can consume, so some remains when the reaction stops. (d) The maximum amount of product predicted by the stoichiometric calculation from the limiting reactant, assuming perfect conversion and no losses. (e) The amount of product actually obtained and isolated in a real experiment. (f) Percent yield = (actual yield ÷ theoretical yield) × 100.

A3. (a) 5.0 × 3/2 = 7.5 mol of O_2. (b) 6.0 × 2/3 = 4.0 mol of $KClO_3$. (c) The $KClO_3$ to KCl ratio is 2 : 2, so 5.0 mol of KCl forms.

A4. (a) 11.0 ÷ 44.01 = 0.250 mol. (b) 0.750 × 58.44 = 43.8 g. (c) 90.0 ÷ 18.02 = 4.99 mol, essentially 5.00 mol. (d) Dividing a mass by the molar mass converts grams into moles; multiplying moles by the molar mass converts back into grams. The molar mass is the bridge in both directions.

A5. (a) Magnesium. It needs 4.0 ÷ 2 = 2.0 mol of O_2, so 4.0 − 2.0 = 2.0 mol of O_2 is left over. (b) Oxygen. It consumes 2 × 1.0 = 2.0 mol of Mg and makes 2.0 mol of MgO, leaving 4.0 − 2.0 = 2.0 mol of Mg. (c) Because the equation may demand the reactants in an unequal ratio. Safety comes only from comparing the amounts through the coefficients, or equivalently from testing how much product each reactant could make.

B1. (a) 10.0 ÷ 16.04 = 0.623 mol of CH_4 (molar mass of CH_4 = 12.01 + 4 × 1.008 = 16.04 g/mol). (b) The ratio is 1 : 1, so 0.623 mol of CO_2, with mass 0.623 × 44.01 = 27.4 g. (c) O_2 consumed = 2 × 0.623 = 1.25 mol, with mass 1.25 × 32.00 = 39.9 g.

B2. (a) Mg: 8.00 ÷ 24.31 = 0.329 mol. O_2: 6.00 ÷ 32.00 = 0.188 mol. (b) From the Mg: 0.329 mol of MgO. From the O_2: 0.188 × 2 = 0.375 mol of MgO. Magnesium produces less, so magnesium is limiting, even though the O_2 had the smaller mass and fewer moles. (c) 0.329 × 40.31 = 13.3 g of MgO. (d) O_2 consumed = 0.329 ÷ 2 = 0.165 mol, so 0.188 − 0.165 = 0.0230 mol remains, with mass 0.0230 × 32.00 = 0.735 g.

B3. (a) 20.4 ÷ 25.5 × 100 = 80.0%. (b) 0.850 × 40.0 = 34.0 g. (c) Any two of: the product is still wet with solvent or water, adding mass that is not product; the product is contaminated with unreacted reactant or a side product; the solid was not fully dried before weighing. A weighing or recording

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error is also possible. A genuine yield of pure product cannot exceed the theoretical maximum. (d) A 90.0% yield needs 0.900 × 25.5 = 22.95 g, so 22.95 − 20.4 = 2.55 g more product.

B4. (a) 0.0250 L × 0.200 = 0.00500 mol of HCl. (b) The ratio is 1 : 1, so 0.00500 mol of NaOH is needed: volume = 0.00500 ÷ 0.150 = 0.0333 L = 33.3 mL. (c) Moles of H_2SO_4 = 0.0200 × 0.500 = 0.0100 mol. The ratio is 1 : 2, so NaOH needed = 2 × 0.0100 = 0.0200 mol. (d) Volume = 0.0200 ÷ 0.400 = 0.0500 L = 50.0 mL. (e) NaCl formed = 0.00500 mol (the 1 : 1 ratio again), with mass 0.00500 × 58.44 = 0.292 g.

B5. (a) N_2: 14.0 ÷ 28.02 = 0.500 mol. H_2: 4.00 ÷ 2.016 = 1.98 mol. (b) Product from the N_2: 2 × 0.500 = 1.00 mol of NH_3. Product from the H_2: 1.98 × 2/3 = 1.32 mol of NH_3. Nitrogen makes less, so N_2 is limiting. (Equivalently, 0.500 mol of N_2 needs 1.50 mol of H_2, and 1.98 mol is available.) (c) 1.00 mol × 17.03 = 17.0 g of NH_3. (d) 12.8 ÷ 17.0 × 100 = 75.2%. (e) H_2 consumed = 1.50 mol, leaving 1.98 − 1.50 = 0.485 mol, with mass 0.485 × 2.016 = 0.978 g.

B6. (a) 13.1 ÷ 65.38 = 0.200 mol of Zn. (b) The ratio is 1 : 2, so HCl consumed = 2 × 0.200 = 0.401 mol. (c) H_2 formed = 0.200 mol, with mass 0.200 × 2.016 = 0.404 g. (d) $ZnCl_2$ formed = 0.200 mol, with mass 0.200 × 136.28 = 27.3 g. Check: reactants used = 13.1 g of Zn + 0.401 × 36.46 = 14.6 g of HCl, totalling 27.7 g. Products = 27.3 + 0.404 = 27.7 g. The totals agree, as conservation of mass requires.

C1. (a) $CaCO_3$: 20.0 ÷ 100.09 = 0.200 mol. HCl: 0.100 L × 3.00 = 0.300 mol. (b) The 0.200 mol of $CaCO_3$ would need 2 × 0.200 = 0.400 mol of HCl, but only 0.300 mol is present, so HCl is limiting. (Equally, product from HCl is 0.150 mol of CO_2 against 0.200 mol from the carbonate.) (c) CO_2 = 0.300 ÷ 2 = 0.150 mol, with mass 0.150 × 44.01 = 6.60 g. (d) 5.94 ÷ 6.60 × 100 = 90.0%. (e) $CaCO_3$ consumed = 0.150 mol, so 0.200 − 0.150 = 0.0498 mol remains, with mass 0.0498 × 100.09 = 4.99 g.

C2. (a) C_3H_8: 11.0 ÷ 44.09 = 0.249 mol (molar mass 3 × 12.01 + 8 × 1.008 = 44.09 g/mol). O_2: 30.0 ÷ 32.00 = 0.938 mol. (b) Burning all the propane needs 5 × 0.249 = 1.25 mol of O_2, but only 0.938 mol is present, so O_2 is limiting. (Equally, the O_2 can burn 0.938 ÷ 5 = 0.188 mol of propane.) (c) CO_2 = 3 × 0.188 = 0.563 mol, with mass 0.563 × 44.01 = 24.8 g. H_2O = 4 × 0.188 = 0.750 mol, with mass 0.750 × 18.02 = 13.5 g. (d) Propane consumed = 0.188 mol, leaving 0.249 − 0.188 = 0.0619 mol, with mass 0.0619 × 44.09 = 2.73 g.

C3. (a) Claim 1 is wrong. Moles present: Mg = 8.00 ÷ 24.31 = 0.329 mol; O_2 = 6.00 ÷ 32.00 = 0.188 mol. MgO possible from the Mg is 0.329 mol; from the O_2 it is 0.188 × 2 = 0.375 mol. Magnesium makes less product, so magnesium is limiting despite its larger mass. Mass alone, without moles and coefficients, cannot identify the limiting reactant. (b) Claim 2 is wrong. The theoretical yield is the maximum possible mass of pure product, so 110% cannot be achieved by any improvement in technique. It means the weighed material was not pure dry product, for example still wet or contaminated, or that the theoretical yield was miscalculated, perhaps from the excess reactant. (c) Claim 3 is wrong. The equation consumes Mg and O_2 in a 2 : 1 mole ratio, not 1 : 1. The 2.0 mol of

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Mg consumes only 1.0 mol of O_2 and makes 2.0 mol of MgO; 1.0 mol of O_2 is left over, so oxygen was in excess even though the starting moles were equal.

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