

Chem 151 Drilling Question Set #5 (Answer key)
Molarity Calculations, Limiting reagents, Percentage Yield

1. Sodium bromate reacts with sodium bromide according to the reaction:



A procedure calls for 5.00 mL of 0.300 M NaBr solution (in 1.0 M H₂SO₄) added to 1.00 mL of 0.250 M NaBrO₃ solution to produce Br₂.

- a) How many moles of each reactant have been mixed together?

Moles of NaBr in solution = (0.300 mol/L)(0.00500 L) = 0.00150 mol NaBr

Moles of NaBrO₃ in solution = (0.250 mol/L)(0.00100 L) = 0.000250 mol NaBrO₃

- b) Which reagent is the limiting reagent of this reaction, NaBr or NaBrO₃?

How many moles of NaBr is needed to react with 0.000250 mol of NaBrO₃ ?

$$(0.000250 \text{ mol NaBrO}_3) \left(\frac{5 \text{ mol NaBr}}{1 \text{ mol NaBrO}_3} \right)$$

$$= 0.00125 \text{ mol NaBr}$$

Limiting reagent is the reagent that limits the amount of product produced from a reaction.

Since there are 0.00150 mol NaBr in the reactant, there is more than enough of NaBr. NaBr is the excess reagent and NaBrO₃ is the limiting reagent.

Alternately, we can calculate the moles of NaBrO₃ needed to react with 0.0015 mol of NaBr.

$$(0.0015 \text{ mol NaBr}) \left(\frac{1 \text{ mol NaBrO}_3}{5 \text{ mol NaBr}} \right) = 0.00030 \text{ mol NaBrO}_3$$

Limiting reagent is the reagent that is insufficient to react with all of the other reagents.

Since there are only 0.00025 mol NaBrO₃, it is not enough to react with all of the NaBr. Hence, NaBrO₃ is the limiting reagent.

- c) How much Br₂ can be produced?

$$(0.00025 \text{ mol NaBrO}_3) \left(\frac{3 \text{ mol Br}_2}{1 \text{ mol NaBrO}_3} \right) = 0.00075 \text{ mol Br}_2$$

Always use the limiting reagent to predict for the amount of product formed from a reaction.

- d) How much excess reagent remains after the reaction?

Moles of NaBr remains = 0.00150 mol – 0.00125 mol = 0.00025 mol NaBr

- e) If the Br₂ produced from reaction reacts with the Br⁻ remains in solution to form Br₃⁻, assuming the reaction goes to completeness, how much Br₃⁻ can be produced?



Since there are 0.00075 mol Br₂, but only 0.00025 mol Br⁻, Br⁻ is the limiting reagent. 0.00025 mol of Br₃⁻ will be produced in the solution.

2. Joan followed the procedures in Question 1 and obtained 0.100 g of Br₂.

a) What was the percentage yield of Joan's procedure?

$$\text{Percentage yield} = \frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100\%$$

Actual yield = 0.100 g

Theoretical yield = 0.00075 mol Br₂ = 0.12 g

$$\text{Percentage yield} = \frac{0.100 \text{ g}}{0.12 \text{ g}} \times 100\% = 83\%$$

1. If Joan extracted all the Br₂ produced into cyclohexane to make 10.0 mL of cyclohexane solution, what is the molarity of Br₂ in the solution?

$$\text{Molarity of Br}_2 = \frac{\text{Moles of Br}_2}{\text{Liter of solution}} = \frac{0.00075 \text{ mol} \times 83\%}{0.0100 \text{ L}} = 0.062 \text{ M}$$

3. 20.0 mL of 0.050 M BaCl₂ solution is mixed with 30.0 mL of 0.050 M Na₂SO₄ solution and BaSO₄ precipitates. Assume the reaction goes to completeness.

a) How many moles of BaSO₄ are formed from the reaction?

$$\text{Moles of Ba}^{2+} = (0.050 \text{ mol/L})(0.0200 \text{ L}) = 0.0010 \text{ mol Ba}^{2+}$$

$$\text{Moles of SO}_4^{2-} = (0.050 \text{ mol/L})(0.0300 \text{ L}) = 0.0015 \text{ mol SO}_4^{2-}$$

Ba²⁺ is the limiting reagent and SO₄²⁻ is the excess reagent. Therefore, 0.0010 mol of BaSO₄ is formed from reaction.

b) What is the molarity of SO₄²⁻ ions in the resultant solution?

$$\text{Moles of unreacted SO}_4^{2-} \text{ in solution} = 0.0015 \text{ mol} - 0.0010 \text{ mol} = 0.0005 \text{ mol}$$

$$\text{Molarity of SO}_4^{2-} \text{ in solution} = \frac{0.0005 \text{ mol}}{(0.0200 + 0.0300) \text{ L solution}} = 0.01 \text{ M}$$

c) What is the molarity of Na⁺ ions in the resultant solution?

Na⁺ is the spectator ion, however, its concentration in the resultant solution will be different because the solution has been diluted to 50.0 mL.

$$\text{Moles of Na}^+ = (0.050 \text{ mol/L})(0.0300 \text{ L}) \left(\frac{2 \text{ mol Na}^+}{1 \text{ mol Na}_2\text{SO}_4} \right) = 0.0030 \text{ mol}$$

$$\text{Molarity of Na}^+ = \frac{0.0030 \text{ mol}}{0.0500 \text{ L solution}} = 0.060 \text{ M}$$

4. Cobalt(II) sulfate can be converted to $\text{Co}(\text{NH}_3)_6\text{Cl}_3$ by a series of reactions. What mass of this compound can be obtained from the complete conversion of 5.30 g of CoSO_4 ? If a student uses 5.30 g of CoSO_4 and produces 5.30 g of $\text{Co}(\text{NH}_3)_6\text{Cl}_3$, what is the percentage yield of this student's synthesis?

Molar masses:
 $\text{CoSO}_4 = 155.0 \text{ g/mol}$
 $\text{Co}(\text{NH}_3)_6\text{Cl}_3 = 267.5 \text{ g/mol}$

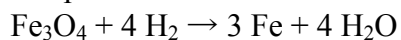
The percentage yield is not 100%.

Theoretical Yield:

$$(5.30 \text{ g } \text{CoSO}_4) \left(\frac{1 \text{ mol } \text{CoSO}_4}{155.0 \text{ g}} \right) \left(\frac{1 \text{ mol } \text{Co}(\text{NH}_3)_6\text{Cl}_3}{1 \text{ mol } \text{CoSO}_4} \right) \left(\frac{267.5 \text{ g}}{1 \text{ mol } \text{Co}(\text{NH}_3)_6\text{Cl}_3} \right) = 9.15 \text{ g } \text{Co}(\text{NH}_3)_6\text{Cl}_3$$

$$\text{Percentage yield} = \frac{5.30 \text{ g}}{9.15 \text{ g}} \times 100\% = 57.9\%$$

5. A sample containing 75% (by mass) Fe_3O_4 reacted with an excess amount of hydrogen gas at high temperature. The water vapor formed from the reaction below was condensed and found to weigh 8.0 g.



- a. How many grams of iron metal were produced from the reaction?

Moles of Fe produced can be calculated from the moles of H_2O formed.

$$\text{Moles of } \text{H}_2\text{O} \text{ formed} = \frac{8.0 \text{ g}}{18.01 \text{ g/mol}} = 0.44 \text{ mol}$$

$$\text{Moles of Fe produced} = (0.44 \text{ mol } \text{H}_2\text{O}) \left(\frac{3 \text{ mol Fe}}{4 \text{ mol } \text{H}_2\text{O}} \right) \left(\frac{55.85 \text{ g}}{1 \text{ mol Fe}} \right) = 18.4 \text{ g Fe}$$

- b. How many grams of the Fe_3O_4 sample had reacted?

Grams of sample reacted =

$$(0.44 \text{ mol } \text{H}_2\text{O}) \left(\frac{1 \text{ mol } \text{Fe}_3\text{O}_4}{4 \text{ mol } \text{H}_2\text{O}} \right) \left(\frac{231.5 \text{ g } \text{Fe}_3\text{O}_4}{1 \text{ mol } \text{Fe}_3\text{O}_4} \right) \left(\frac{100 \text{ g sample}}{75 \text{ g } \text{Fe}_3\text{O}_4} \right) = 34 \text{ g sample}$$

6. Consider the reaction: $2 \text{Al} + 3 \text{Br}_2 \rightarrow 2 \text{AlBr}_3$

- a) For a reaction mixture that contains 10.0 g of Al and 20.0 g of Br_2 , if the reaction goes to completeness, how many grams of AlBr_3 is formed?

$$\text{Moles of Al} = \frac{10.0 \text{ g}}{26.98 \text{ g/mol}} = 0.371 \text{ mol Al}$$

$$\text{Moles of Br}_2 = \frac{20.0 \text{ g}}{159.8 \text{ g/mol}} = 0.125 \text{ mol Br}_2$$

Moles of Al required to react with 0.125 mol Br_2

$$= (0.125 \text{ mol Br}_2) \left(\frac{2 \text{ mol Al}}{3 \text{ mol Br}_2} \right) = 0.0833 \text{ mol Al}$$

Since there are 0.371 mol Al, Al is the excess reagent and Br_2 is the limiting reagent.

$$\text{Grams of AlBr}_3 \text{ formed} = (0.0833 \text{ mol Al}) \left(\frac{2 \text{ mol AlBr}_3}{2 \text{ mol Al}} \right) \left(\frac{266.7 \text{ g}}{1 \text{ mol AlBr}_3} \right) = 22.2 \text{ g}$$

- b) How many grams of excess reagent remains after the reaction?

$$\text{Moles of excess Al} = 0.371 \text{ mol} - 0.0833 \text{ mol} = 0.288 \text{ mol}$$

$$\text{Grams of excess Al} = (0.288 \text{ mol})(26.98 \text{ g/mol}) = 7.76 \text{ g}$$