

Chemistry Packet

**Day 1**

Complete the "Balancing and Stoichiometry Review" worksheet

**Day 2**

Review the "Limiting Reagents, Theoretical, Actual and Percent Yields" notes

\*email me if you have questions, [nicole.hoover@k12.wv.us](mailto:nicole.hoover@k12.wv.us)

**Day 3**

Complete the "Stoichiometry and Limiting Reactants" worksheet

**Day 4**

Review the "Excess Reactant" notes

\*email me if you have questions

**Day 5**

Complete all parts of number 1 on the "Limiting Reactants" worksheet

**Day 6**

Complete all parts of number 2 on the "Limiting Reactants" worksheet

**Day 7**

Complete all parts of number 1 on the "Limiting Reactants, Theoretical Yield and Percent Yield" worksheet

**Day 8**

Complete all parts of number 2 on the "Limiting Reactants, Theoretical Yield and Percent Yield" worksheet

**Day 9**

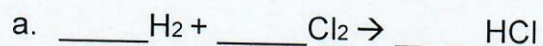
Complete all parts of number 3 on the "Limiting Reactants, Theoretical Yield and Percent Yield" worksheet

**Day 10**

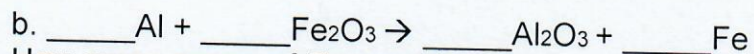
Complete all parts of number 4 on the "Limiting Reactants, Theoretical Yield and Percent Yield" worksheet

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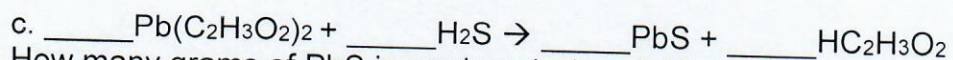
### Balancing and Stoichiometry Review



How many grams of  $\text{HCl}$  can be produced if 7.25 g of  $\text{Cl}_2$  is reacted with an unlimited supply of  $\text{H}_2$ ?



How many grams of  $\text{Fe}$  can be produced when 10.0g of  $\text{Al}$  is reacted with an excess (unlimited) supply of  $\text{Fe}_2\text{O}_3$ ?



How many grams of  $\text{PbS}$  is produced when 5.00g of  $\text{H}_2\text{S}$  is reacted with an excess (unlimited) supply of  $\text{Pb}(\text{C}_2\text{H}_3\text{O}_2)_2$ ?

## LIMITING REAGENTS, THEORETICAL , ACTUAL AND PERCENT YIELDS

A limiting reagent is a chemical *reactant* that limits the amount of product that is formed. The limiting reagent gives the smallest yield of product calculated from the reagents (reactants) available. This smallest yield of product is called the theoretical yield.

To find the limiting reagent and theoretical yield, carry out the following procedure:

1. Find the moles of each reactant present.
2. Calculate the moles of a product formed from each mole of reactant.
3. Identify the reactant giving the smaller number of moles of product. This reactant is the Limiting Reagent:
4. Calculate the grams of product produced by the Limiting Reagent. This is the theoretical yield.

### THEORETICAL YIELD

The theoretical yield is the amount of the product in g formed from the limiting reagent. From the moles of limiting reagent available, calculate the grams of product that is theoretically possible (same as Step 4 above).

### ACTUAL YIELD

The actual yield is the amount of the product in g actually formed in the laboratory.

### PERCENT YIELD

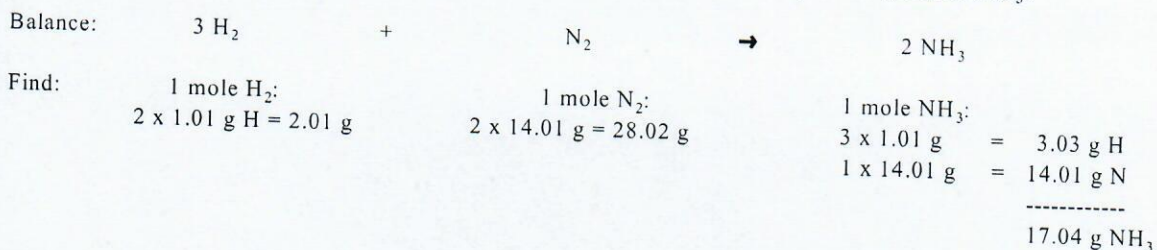
The percent yield is the percent of the product formed based upon the theoretical yield.

$$\frac{\text{actual yield in g}}{\text{theoretical yield in g}} \times 100 \% = \text{Percent Yield}$$

## LIMITING REAGENTS, THEORETICAL, ACTUAL AND PERCENT YIELDS

### EXAMPLE OF A LIMITING REAGENT PROBLEM

How many grams of  $\text{NH}_3$  can be produced (theoretically) from the reaction of 5.0 g of  $\text{N}_2$  and 5.0 g  $\text{H}_2$ ? What is the limiting reagent? If 8.52 g are actually formed, what is the percent yield of  $\text{NH}_3$ ?



To answer this type of question, carry out EACH step of the following procedure:

- Find the g of product produced from each of the given amounts of reactants.

$$\text{For H}_2: 5.0 \text{ g H}_2 \times \frac{1 \text{ mole H}_2}{2.02 \text{ g H}_2} \times \frac{2 \text{ mole NH}_3}{3 \text{ mol H}_2} \times \frac{17.04 \text{ g NH}_3}{1 \text{ mol NH}_3} = 28.12 \text{ g NH}_3$$

$$\text{For N}_2: 5.0 \text{ g N}_2 \times \frac{1 \text{ mole N}_2}{140.01 \text{ g N}_2} \times \frac{2 \text{ mole NH}_3}{1 \text{ mol N}_2} \times \frac{17.04 \text{ g NH}_3}{1 \text{ mol NH}_3} = 12.16 \text{ g NH}_3$$

(carry 1 or 2 additional significant figures here)

- Identify the smaller amount of calculated product:

This is the maximum amount of product that can be formed. The larger amount cannot be formed because there is not enough of the limiting reagent present.

12.16 g is the smaller amount calculated.

- Identify the reactant that gives the smallest amount of product.  
*The limiting reagent* determines the amount of product formed.

The limiting reagent is  $\text{N}_2$ .

- Identify the **THEORETICAL YIELD**:

The theoretical yield is the amount of the product (in g) formed from the limiting reagent.

12 g is the theoretical yield

- Identify the **ACTUAL YIELD**:

The actual yield is the amount of the product (in g or mol) actually formed in the laboratory.

8.25 g is the actual yield.

- Calculate the **PERCENT YIELD**: The percent yield is based upon the theoretical yield.

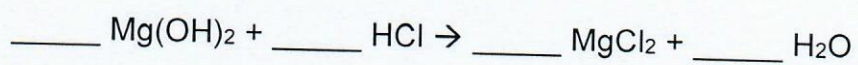
$$\frac{\text{actual yield (g)}}{\text{theoretical yield (g)}} \times 100\% = \text{Percent Yield} = \frac{8.25 \text{ g}}{12.16 \text{ g}} \times 100\% = 68\%$$

(For theoretical yield, carry more sig. figs. in this calculation)

Final sig. figs. limited by the given amounts of reactants.

## Stoichiometry and Limiting Reactants

1. Balance the following equation:



2. How many grams of magnesium hydroxide would be required to produce 250 grams of water?
3. If you performed the reaction in problem 1 using 100 grams of hydrochloric acid and 55 grams of magnesium hydroxide, how many grams of magnesium chloride would you expect to make (theoretical yield)?
4. What is the limiting reactant in this reaction?

## Excess Reactants

Given the chemical equation and the masses of reactants, determine the mass of excess reactant.

### Example:

A 2.00 g sample of ammonia reacts with 4.00 g of oxygen according to the equation  
 $4 \text{NH}_3 + 5 \text{O}_2 \rightarrow 4 \text{NO} + 6 \text{H}_2\text{O}$ .

How much excess reactant remains after the reaction has stopped?

### Strategy

1. Balance the equation
2. Calculate the moles of product from the first reactant
3. Calculate the moles of product from the second reactant
4. Identify the limiting reactant (the one that produces less product)
5. Calculate the mass of the excess reactant used
6. Calculate the mass of the unused excess reactant

### Moles of NO from NH<sub>3</sub>

Convert grams of NH<sub>3</sub> to moles of NH<sub>3</sub>, and then use the molar ratio from the equation to get moles of NO.

The molar mass of NH<sub>3</sub> is 17.03 g/mol.

Moles of NH<sub>3</sub> =  $2.00 \text{g NH}_3 \times 1 \text{ mol NH}_3 / 17.03 \text{g NH}_3 = 0.1174 \text{ mol NH}_3$

$0.1174 \text{ mol NH}_3 \times 4 \text{ mol NO} / 4 \text{ mol NH}_3 = 0.1174 \text{ mol NO}$

### 3. Moles of NO from O<sub>2</sub>

The molar mass of O<sub>2</sub> is 32.00 g/mol.

Moles of O<sub>2</sub> =  $4.00 \text{g O}_2 \times 1 \text{ mol O}_2 / 32.00 \text{g O}_2 = 0.1250 \text{ mol O}_2$

$0.1250 \text{ mol O}_2 \times 4 \text{ mol NO} / 5 \text{ mol O}_2 = 0.1000 \text{ mol NO}$

### 4. Identify limiting and excess reactants

O<sub>2</sub> is the limiting reactant, since it gives the smaller amount of NO.

NH<sub>3</sub> is the only other reactant, so it is the excess reactant.

### 5. Calculate the mass of excess reactant used up.

Use the molar ratio from the equation to convert moles of O<sub>2</sub> (from Step 3) to moles of NH<sub>3</sub>, and then convert moles of NH<sub>3</sub> to grams of NH<sub>3</sub>.

$0.1250 \text{ mol O}_2 \times 4 \text{ mol NH}_3 / 5 \text{ mol O}_2 = 0.1000 \text{ mol NH}_3$

$0.1000 \text{ mol NH}_3 \times 17.03 \text{ g NH}_3 / 1 \text{ mol NH}_3 = 1.703 \text{ g NH}_3$

### 6. Calculate the mass of unused excess reactant.

We started with 2.00 g of NH<sub>3</sub> and used up 1.703 g, so

Mass of excess NH<sub>3</sub> =  $2.00 \text{ g} - 1.703 \text{ g} = 0.30 \text{ g NH}_3$

## Limiting Reactants

1. Take the reaction:  $4 \text{NH}_3 + 5 \text{O}_2 \rightarrow 4 \text{NO} + 6 \text{H}_2\text{O}$ . In an experiment, 3.25 g of  $\text{NH}_3$  are allowed to react with 3.50 g of  $\text{O}_2$ .

a. What is the theoretical yield?

b. Which reactant is the limiting reagent?

c. How much of the excess reactant remains after the reaction?

2. Consider the reaction of  $\text{C}_6\text{H}_6 + \text{Br}_2 \rightarrow \text{C}_6\text{H}_5\text{Br} + \text{HBr}$ . In an experiment, 42.1 g of  $\text{C}_6\text{H}_6$  react with 73.0 g of  $\text{Br}_2$ .

a. What is the theoretical yield?

b. What is the limiting reagent?

c. How much of the excess reactant remains after the reaction?

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### Limiting Reactants, Theoretical Yield and Percent Yield

1. Consider the following reaction:  $2 \text{ Al} + 6 \text{ HBr} \rightarrow 2 \text{ AlBr}_3 + 3 \text{ H}_2$
- When 3.22 moles of Al react with 4.96 moles of HBr, how many grams of  $\text{H}_2$  are formed (theoretical yield)?
  - What is the limiting reactant?
  - For the reactant in excess, how many grams are left over at the end of the reaction?
  - If the experimental yield of the reaction was 3.78 grams of  $\text{H}_2$ , what was the percent yield?



2. Consider the following reaction:  $3 \text{ Si} + 2 \text{ N}_2 \rightarrow \text{Si}_3\text{N}_4$
- When 21.44 moles of Si react with 17.62 moles of  $\text{N}_2$ , how many grams of  $\text{Si}_3\text{N}_4$  are formed (theoretical yield)?
  - What is the limiting reactant?
  - For the reactant in excess, how many grams are left over at the end of the reaction?
  - If the experimental yield of the reaction was 992.6 grams of  $\text{Si}_3\text{N}_4$ , what was the percent yield?

3. Consider the following reaction:  $2 \text{CuCl}_2 + 4 \text{KI} \rightarrow 2 \text{CuI} + 4 \text{KCl} + \text{I}_2$
- a. When 0.56 moles of  $\text{CuCl}_2$  react with 0.64 moles of  $\text{KI}$ , how many grams of  $\text{I}_2$  are formed (theoretical yield)?

b. What is the limiting reactant?

c. For the reactant in excess, how many grams are left over at the end of the reaction?

d. If the experimental yield of the reaction was 38.6 grams of  $\text{I}_2$ , what was the percent yield?

4. Consider the following reaction:  $4 \text{FeS}_2 + 11 \text{O}_2 \rightarrow 2 \text{Fe}_2\text{O}_3 + 8 \text{SO}_2$
- a. When 26.62 moles of  $\text{FeS}_2$  react with 5.44 moles of  $\text{O}_2$ , how many grams of  $\text{SO}_2$  are formed (theoretical yield)?

b. What is the limiting reactant?

c. For the reactant in excess, how many grams are left over at the end of the reaction?

d. If the experimental yield of the reaction was 234.7 grams of  $\text{I}_2$ , what was the percent yield?