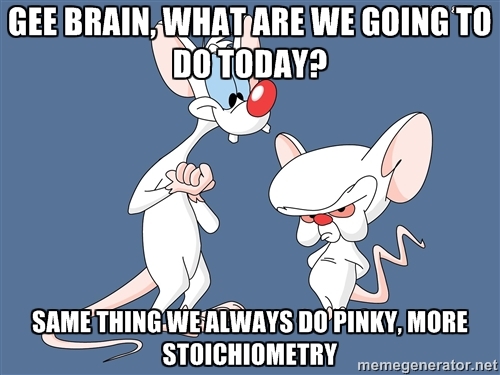
*Unit 9: Stoichiometry*

*Stoy-kee-om-uh-tree.*



*Important Vocabulary and formulas for Unit 9*

Mole ratio

Limiting Reagent

Excess Reagent

Theoretical Yield

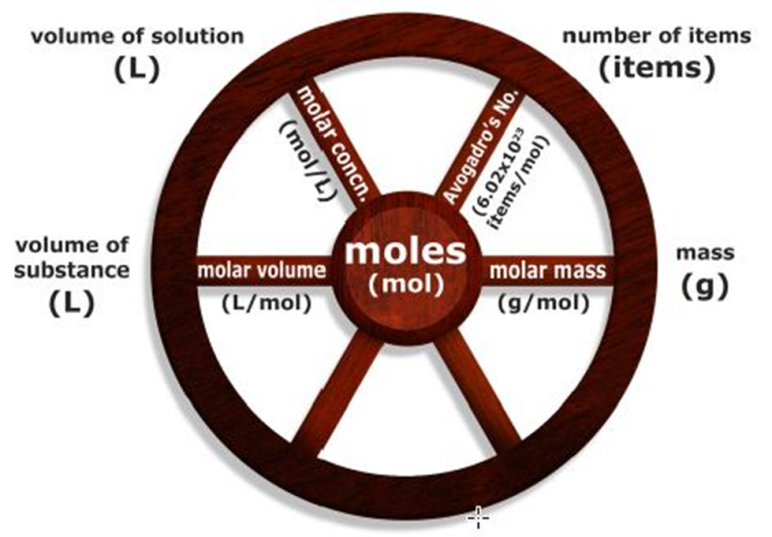
Actual Yield

Percent Yield

Hydrate

Empirical Formula

Molecular Formula



*Notes: A review of the Mole*

The most important skill for you to have mastered BEFORE we begin the stoichiometry unit is to be able to convert the mass of a compound into moles and the number of moles of a compound into mass. Please review accordingly.

1. Express 3.50 moles potassium sulfate K2SO4 in grams. Calculate the percent composition of potassium.
2. Convert 75.0 grams carbon disulfide CS2 to moles. Calculate the percent composition.

The Unique Thing about Gases…

The amount of space that 1 mole of a gas takes up is roughly the same as any other gas under the same conditions. We’ll learn why later but for now what you need to know is that at standard temperature and pressure (STP) one mole of any gas has a volume of 22.4 Liters. Standard temperature is 0 0C and standard pressure is 1 atm which stands for atmosphere. That’s not relevant now but it will be next unit.

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**Empirical and Molecular Formulas**

Once a scientist has determined the percent composition of the various elements in the sample the empirical formula can then be determined. ***The empirical formula represents the smallest whole-number ratio of the elements in the compound.*** When determining the empirical formula always assume that you have a 100-gram sample; therefore the percent composition of each element can be used as the number of grams that you have of each element.

There are three steps in determining the empirical formula of a compound:

Step 1: Change the number of grams given to moles by using the atomic mass of each element. Use at least 4 significant figures!

Step 2: Determine the mole ratio by dividing each number of moles by the smallest number of moles. If this does produce whole number ratios of moles, multiply all mole ratios by the smallest whole number so that you do have whole numbers.

Step 3: Use the ratio as subscripts in the formula. The elements usually appear in the problem in the same order that they should appear in the formula.

Example 1: A compound is 12.69% aluminum, 19.73% nitrogen, and 67.57% oxygen. What is the empirical formula of the sample?

Example 2: A 250.0 g sample of a compound contains 171.0 grams of chromium and the remainder is oxygen. Calculate the empirical formula of this compound.

Since the empirical formula only gives the simplest whole number ratio of atoms, it is sometimes necessary to determine what the molecular formula for the compound would be. For example the empirical formula for hydrogen peroxide is HO, but the correct molecular formula is H2O2. In order to determine the molecular formula both the empirical formula and the molecular (formula) mass must be known.

In general: **molecular formula = (empirical formula) × n**

To determine ‘n’ you must first determine the empirical formula’s mass. Once this is done a simple calculation will give you the value of ‘n’:



Once the value of ‘n’ is known simple multiply the subscripts in the empirical formula by ‘n’ to obtain the correct subscripts for the molecular formula.

Example 1: In a 50.0 gram sample of hydrazine (a chemical used to treat waste water) there are found to be 43.8 grams of nitrogen and the remainder is hydrogen. The molecular mass of hydrazine is 32.0 grams per mole. What is the molecular formula of hydrazine?

Example 2: A compound is composed of 7.20 grams of carbon, 1.20 grams of hydrogen, and 9.60 grams of oxygen. The molecular mass of the compound is 180 grams. What are the empirical and molecular formulas for this compound?

*Notes: Stoichiometry*

In this unit, you will be asked to calculate the masses of reagents and products based on a specific chemical recipe – the balanced equation.

You will need to be able to convert masses to moles and moles to masses. To convert from mass to moles, you will divide the given mass by the molar mass. To convert from moles to mass, you will multiply the given number of moles by the molar mass. Review your mole map. Also, mass means grams!

The new idea is the concept of the MOLE RATIO. A mole ratio can be created for any two substances in a balanced equation. The numbers for the mole ratio come from the COEFFICIENTS in the balanced equation.

Example #1: How many **moles of water** can be made when **6 moles of oxygen** gas react with excess hydrogen?

First we need a balanced equation: 2 H2 + O2 🡪 2 H2O

According to this equation, **2 moles** of hydrogen react with **1 mole** of oxygen to produce **2 moles** of water. Do you see where the numbers came from? Can you write a ratio (fraction) using these numbers that relates oxygen and water?

Let’s write a mole ratio that relates oxygen to water. There are two ways to do this:

or .

These are really **conversion factors** in disguise! You know what’s coming…

Finally, use dimensional analysis to solve! Remember to start with the number that you are given in the question, then arrange the conversion factors in such a way that the units you don’t want cancel out.



6 moles oxygen x = 12 moles of water.



Draw a line through the units that cancel. Do you understand how we got 12 moles of water? If so, you understand stoichiometry. The rest is just details!



Example #2: Calculate the number of moles of carbon dioxide that are formed when 10 moles of methane (CH4) combusts.

Here’s another example that shows how stoichiometry is used in the lab.

Example #3: How many **grams of water** can be produced when **6 grams of oxygen** react with excess hydrogen?

Do you see the difference between this problem and the first? What is the difference?

To solve, start off with a balanced equation: 2 H2 + O2 🡪 2 H2O

We will also use a mole ratio that relates what we’re given (oxygen) to what we’re solving for (water):

or . Which one do you think we’ll use? Why?

Here’s the new part: You can only use a mole ratio when you have moles…we don’t have moles, we have grams. Let’s convert grams of oxygen to moles of oxygen.

6 grams O2 x =

That should do it right? But don’t calculate yet! We can do all of the stoichiometry steps at one time!

6 grams O2 x x =



Do you see what we did? What is the last step?



Remember that we’re solving for grams of water. If we stopped here, we would have moles of water so…

6 grams O2 x x x =

Solve this without a calculator! If you need to leave your answer as a fraction then so be it! But make it a proper fraction!

Example #4: How many grams of water will be produced when 12 grams of hydrogen react? How many grams of oxygen gas are required to react with the 12 grams of hydrogen? Do you notice anything special about these numbers?

*Notes: Limiting Reagent*

You may have noticed the word “excess” in several of the previous problems we have worked. What does it mean?! The opposite of excess is “limiting.” What happens when you don’t know which reactant is in excess? Sadly, what it means for you is more work☹. But it’s the same type of work…nothing new. Just more of the same☺

Figure this next problem out and you’ll know all there is to know about the concept of “Limiting Reagent.”

Example #1 How many pancakes can be made with the following recipe and ingredients? Before you get started you may find it helpful to write a “chemical equation” for the recipe.

Recipe:

2 cups flour 2 cups buttermilk

2 tsp. baking powder 3 eggs

1 tsp. baking soda 4 tbsp. butter

½ tsp salt 3 tbsp. Sugar

Yield: 12 pancakes

After searching your pantry and refrigerator, you find that you have 16 cups of flour, 48 tsp. baking powder, 22 tsp. baking soda, 500 tsp. salt, 8 cups of buttermilk, 11 eggs, 8 tbsp. butter and 16 tbsp. sugar.

What ingredient did you base your answer on? Why?!

How many pancakes can you make with 16 cups of flour and excess everything else?

Do you see the difference between this question and the previous? What is the difference?

Example #2 How many moles of water can be made from 21 moles of hydrogen and 10 moles of oxygen? Remember you must start off with a balanced equation!

Example #3 How many grams of water can be made from 10 grams of hydrogen and 64 grams of oxygen? What is the limiting reagent?

*Notes: Percent Yield*

**Theoretical yield** - the maximum amount of product that can be produced from a given amount of reactant. This is the calculated amount of product you should get if everything is perfect.

**Actual yield** - measured amount of product obtained from a reaction. This is amount of product actually produced. Therefore it must be given in the problem or calculated from the percent yield and theoretical yield.

The efficiency of a reaction is measured by comparing the actual yield to theoretical yield. This quantity is called the percent yield and is calculated as follows:

|  |  |  |
| --- | --- | --- |
| Percent yield = | Actual yield | X 100% |
| Theoretical yield |

In all problems concerning percent yield you will be given either the actual yield or the percent yield. You must calculate the theoretical yield as we have done in the past and then use that to calculate what the problem is asking for. Remember, we are talking about what has been produced or yielded so **yields are always products!**

Example Problem 1: A chemist produced only 6.58 g of a compound although her calculations indicated she should have produced 9.00. What was her percent yield?

Example Problem 2: Methanol can be produced through the reaction of CO and H2 in the presence of a catalyst. CO + 2H2 → CH3OH

If 75.0 g of CO reacts to produce 68.4 g of CH3OH, what is the percent yield of CH3OH?

Example Problem 3: Aluminum reacts with excess copper (II) sulfate. If 1.85 g of Al react and the percent yield of copper is 56.6, what mass of Cu was produced?

2 Al + 3 CuSO4 → Al2(SO4)3 + 3 Cu

**STOICHIOMETRIC CALUCLATIONS SUMMARY**

1) Write out reactants and products—use only the charges of each atom to put together the appropriate molecules. Use prefixes for molecular compounds.

2) Balance the Equation!

3) Only mol-mol or mol-kJ conversions using ratios from a balanced equation will work! If necessary,

Convert any masses (grams) to moles by dividing by the molar mass (grams/mol) of the substance given, OR

Convert any volumes (Liters) to moles by dividing by 22.4 (liters/mol).

4) Multiply by the molar ratio (placing the coefficient—from the BALANCED equation—of the moles or kJ you are given in the denominator and the coefficient of the substance or kJ desired in the numerator).

5) If more than 2 reactants are given, repeat steps 3 and 4 for all reactants, then determine the limiting reactant (smallest molar quantity between reactants).

**USE ONLY THE LIMITING REACTANT TO CONTINUE TO STEP 6!**

6) If necessary,

Convert any moles to masses (grams) by multiplying by the molar mass (grams/mol) of the substance desired, OR

Convert any moles to volumes by multiplying by 22.4 (liters/mol).

**Notes on % Yield:**

Only products (quantities that apply to the right side of a balanced equation) can be placed into the Yield Formula, and both actual and theoretical yields must have the same units (and molecule!) If reactants are given, you'll probably be using them to compute a theoretical yield. All calculated quantities are theoretical yields.

% Yield = Actual Yield x 100%

Theoretical Yield

*Notes: Hydrates*

A hydrate is an ionic compound that has water molecules integrated into its crystal structure.

Here is an example: MgSO4 • 7 H2O. The seven water molecules are included in the formula because there are seven molecules for every one magnesium ion and one sulfate ion integrated into the crystalline structure of this salt.

The name MgSO4 • 7 H2O is magnesium sulfate *hepta*hydrate. The same prefixes used for covalent nomenclature are used here to indicate the number of water molecules.

Name the following hydrates

CuSO4 • 5 H2O BaCl­2 • 2 H2O

Na2CO3 • 2 H2O Na2CO3 • 10 H2O

Empirical Formulae of Hydrates

Example #1: A chemist finds by experimentation that a certain compound contains 76.9% dry calcium sulfite and 23.1% water of hydration (by mass) in its crystal. What is the formula of the hydrate?

Example #2: A chemist has a sample of hydrated Li2SiF6­ and it weighs 0.4813 grams. He heats it strongly to drive off the water of hydration, and after heating, he finds that the anhydrous compound has a mass of 0.391 grams. Find the formula of the hydrate.

***Problem Set #1 Empirical and Molecular Formula Homework***

***Find the empirical formula for each of the following from the given percent compositions. Show all work with units.***

1. 30.43% nitrogen, 69.57% oxygen
2. 42.50 % chromium, 57.50% chlorine
3. 12.5% hydrogen, 37.5% carbon, 50.0% oxygen
4. 29.40% calcium, 23.56% sulfur, 47.04 % oxygen

***Find the empirical formula for each of the following. Show all work with units.***

1. A pure sample of a mercury oxide produced 20.3 g of mercury and 1.7 g of oxygen. What is the empirical formula?
2. 0.916 g of iron is heat in air and reacts with the oxygen present. The resulting product weighs 1.178 g. What is the empirical formula?
3. A sample containing a potassium salt has the following composition: 17.96 g of potassium, 7.35 g of sulfur, and 14.70 g of oxygen. What is the empirical formula?
4. Analysis of 100.0 g of a compound produced the following results: 26.6 of potassium, 35.4 g of chromium and the remainder is oxygen. Calculate the empirical formula.

***Calculate the molecular formula for each of the questions below. Show all work with units.***

1. A compound has the following percent composition: 26.7 % carbon, 2.2 % hydrogen, 71.1% oxygen. The molecular mass of this compound is 90.0 g/mole. What is the molecular formula?
2. The percent composition of ethane gas is 80.0% carbon and 20.0% hydrogen. The molecular mass of ethane is 30 g/mole. Find the molecular formula.
3. An unknown compound is analyzed and found to consist of 49.0% carbon, 2.7 % hydrogen and 48.2 % chlorine. Additional experimental data suggests that the molar mass is about 162 g/mole. What is the empirical formula and molecular formula?
4. A compound has the following percent composition: carbon = 40.0%, hydrogen = 6.67% and oxygen = 53.3%. The molecular mass is 120. g/mole. Find the empirical formula.
5. A compound consisting of 82.66% carbon and 17.34% hydrogen has a molecular mass of 58.1 g/mole. What is the molecular formula?
6. When 10.0 g of phosphorous reacted with oxygen, it produced 17.77 g of a phosphorous oxide. This new compound was found to have a formula mass of approximately 220 g/mole. Find the molecular formula.
7. A compound known to contain only hydrogen, carbon and oxygen was subjected to analysis. The composition was found to be 69.96% carbon, 7.83% hydrogen and the remainder was oxygen. The molecular mass of this compound was found to be approximately 1045 g/mole. Find the empirical and molecular formula.

*Problem Set #2: Mole-mole stoichiometry*

1. How many moles of oxygen are produced by the decomposition of 6.25 moles of potassium chlorate?

2 KClO3 → 2 KCl + 3 O2

2. How many moles of hydrogen are produced from the reaction of 3.86 moles of zinc with an excess of hydrochloric acid?

3. How many moles of oxygen are necessary to react completely with 4.50 moles of propane, C3H8?

C3H8 + 5 O2 → 3 CO2 + 4 H2O

4. How many moles of potassium nitrate are produced when 2.15 moles of potassium phosphate react with excess aluminum nitrate?

5. How many moles of sodium chloride are produced when 7.23 moles of iron (II) chloride reacts with excess sodium phosphate? 3 FeCl2 + 2 Na3PO4 → Fe3(PO4)2 + 6 NaCl

6. How many moles of potassium are required to react with 2.50 moles of aluminum chloride to produce potassium chloride?

7. How many moles of oxygen are required to react with 2.50 moles of iron metal to produce iron (III) oxide?

4 Fe + 3O2 → 2 Fe2O3

8. How many moles of hydrofluoric acid are required to produce 5.20 moles of silicon tetrafluoride according to the following reaction? SiO2 + 4 HF → 2 H2O + SiF4

9. How many moles of water are required to produce 6.75 moles of arsenic acid in the reaction with diarsenic pentoxide? As2O5 + 3 H2O → 2 H3AsO4

1. How many moles of carbon dioxide are produced in the combustion of 17.0 moles of octane, C8H18?

*Problem Set #3: practical stoichiometry*

Assume all gases are at STP.

1. How many grams of potassium chloride are produced if 25.0 g of potassium chlorate decompose?

**2 KClO3 → 2 KCl + 3 O2**

2. What volume of ammonia at STP is produced if 30.0 g of nitrogen is reacted with an excess of hydrogen?

**N2 + 3 H2 →2 NH3**

3. What volume of carbon dioxide gas is released when 5.00 L of ethanol, C2H5OH, is burned?

**C2H5OH + 3 O2 → 2 CO2 + 3 H2O**

4. What volume of chlorine is produced when 45.0 g of aluminum chloride is decomposed?

**2 AlCl3 → 2 Al + 3 Cl2**

5. What volume of nitrogen is produced when 18.5 liters of hydrogen reacts with nitrogen monoxide?

**2 H2 + 2 NO → 2 H2O + N2**

6.What mass of potassium sulfate is required to react with 35.0 g of aluminum bromide?

**2 AlBr3 + 3 K2SO4 → 6 KBr + Al2(SO4)3**

7. If 20.0 liters of oxygen are consumed in the reaction below, how many liters of water vapor are produced?

**C3H8 + 5 O2 → 3 CO2 + 4 H2O**

8. What volume of hydrogen at STP is produced when 2.50 g of zinc reacts with an excess of hydrochloric acid?

**Zn + 2 HCl → ZnCl2 + H2**

9. What mass of iron must be used to produce 125 g of iron (III) oxide? **4 Fe + 3 O2 → 2 Fe2O3**

10. What mass of copper metal will be produced when 45.0 L of ammonia gas reacts with copper (II) oxide?

**3 CuO + 2 NH3 → 3 Cu + 3 H2O + N2**

*Problem Set #4: Mixed stoichiometry*

1. How many moles of NO2 are produced by the reaction of 0.772 moles N2 with O2?

2. How many grams of Na2SO4 can be produced by the reaction of 250 g of NaOH with sufficient H2SO4?

3. How many moles of H2O will react with 1.75 moles PCl3 to form HCl and H3PO3?

4. How many moles of FeCl3 will be produced by the reaction of 3.00 mole HCl. Use the equation below.

KMnO4 + 5 FeCl2 + 8 HCI → MnCl2 + 5 FeCl3 + 4 H2O + KCl

5. How many grams of barium hydroxide will be used up in the reaction with hydrochloric acid to produce 45.0 g of barium chloride plus some water?

6. Sulfur trioxide reacts violently with water to produce sulfuric acid. How many grams of sulfur trioxide are necessary to produce 3.21 kg of sulfuric acid?

7. Determine the number of grams of chloric acid that will just reaction with 30.0 g of calcium carbonate to produce carbon dioxide, water, and calcium chlorate.

8. Nitrogen trichloride reacts with water to produce ammonia (NH3) and hypochlorous acid (HClO). Calculate the number of grams of ammonia that can be produced from 20.0 g of nitrogen trichloride.

9. Determine the number of grams of lithium oxide necessary to prepare 53.0 g of lithium hydroxide by addition of excess water.

10. How many grams of barium hydroxide will it take to exactly react with 80.0 g of phosphoric acid to produce barium phosphate and water?

11. How much silver chloride can be produced in the double replacement reaction between 50.0 g barium chloride and silver nitrate?

12. What mass of zinc chloride can be prepared by treating 123g of hydrochloric acid in water solution with excess zinc? The other product is hydrogen gas.

13. What mass (in kg) of aluminum oxide are needed to prepare 5000. g of Al in the reaction below?

Al2O3 + 3 C → 2 Al + 3 CO

14. How many grams of sodium chloride can be produced from 40.0 g of chlorine and excess sodium?

15. How much potassium chlorate must be decomposed to form 3.22 g of oxygen? The other product is potassium chloride.

*Problem Set #5: Limiting Reactant stoichiometry*

1. When 7.24 g of solid magnesium and 3.86 g of oxygen gas react to form magnesium oxide, how many grams of magnesium oxide will be formed? Also, calculate the mass of excess reactant. 2 Mg +O2 → 2 MgO

2. If 15.5 grams of aluminum are reacted with 46.7 grams of chlorine gas, how many grams of aluminum chloride are formed? Also, calculate the mass of excess reactant. 2 Al + 3 Cl2 → 2 AlCl3

3. What mass of calcium nitrate can be prepared by the reaction of 18.9 grams of nitric acid with 7.40 grams of calcium hydroxide? Also, calculate the mass of excess reactant. 2 HNO3 + Ca(OH)2 → Ca(NO3)2 + 2 H2O

4. If 15.5 g of lead (II) nitrate are reacted with 3.81 g of sodium chloride, what mass of lead (II) chloride is formed? Also, calculate the mass of excess reactant.

Pb(NO3)2 + 2 NaCl → PbCl2 + 2 NaNO3

5. If 8.45 g of aluminum metal are reacted with 76.5 g of copper (II) sulfate, find the mass of copper formed. Also, calculate the mass of excess reactant.

2 Al + 3CuSO4 → Al2(SO4)3 + 3 Cu

6. How many grams of aluminum sulfide can form from the reaction of 9.00 g of aluminum and 8.00 g of sulfur? Also, calculate the mass of excess reactant.

2Al + 3S 🡪 Al2S3

7. 75.0 g of zinc are added to 125 g of nitric acid. What volume of hydrogen gas (at STP) is released? Also, calculate the mass of excess reactant.

Zn + 2HNO3 🡪 Zn(NO3)2 + H2

8. 11.0 L of hydrogen and 5,00 L of oxygen are exploded together in a reaction tube. What volume of water vapor was formed? Also, calculate the mass of excess reactant. 2H2 + O2 🡪 2H2O

9. How many grams of iron (III) iodide are formed when 25.7 g of iron reacts with 115 grams of iodine? Also, calculate the mass of excess reactant.

2 Fe +3 I2 → FeI3

10. If 75,0 g of potassium hydroxide are permitted to react with 135 g of sulfuric acid, what mass of potassium sulfate is formed? Also, calculate the mass of excess reactant. 2 KOH + H2SO4 → K2SO4 + 2 H2O

*Problem Set #6: Stoichiometry with percent yield.*

**Show all work - don't forget units!! Assume all gases are measured at STP.**

1. If 12.5 g of copper are reacted with an excess of chlorine gas, 25.4 g of copper (II) chloride are obtained. What is the percent yield? Cu + Cl2 → CuCl2

2. A chemistry student uses 25.0 g of lithium sulfate to react with excess barium nitrate. The student isolates 43.5 g of barium sulfate. Find the percent yield of this reaction.

Li2SO4 + Ba(NO3)2 → BaSO4 + 2LiNO3

3. What is the percent yield if 5.50 g of hydrogen react with nitrogen to form 20.4 g of ammonia?

N2 + 3 H2 → 2 NH3

4. What is the percent yield if 3.74 g of copper is produced when 1.87 g aluminum is reacted with an excess of copper (II) sulfate? 2 Al + 3 CuSO4 → Al2(SO4)3 + 3 Cu

5. When 50.0 g of silicon dioxide is heated with an excess of carbon, 32.2 g of silicon carbide is produced. Carbon monoxide is the other product. What is the percent yield of this reaction?

SiO2 + 4 C → SiC2 + 2 CO

6. When 84.8 g of iron (III) oxide reacts with an excess of carbon monoxide, then 57.3 g of iron is produced. Carbon dioxide is the other product. What is the percent yield of this reaction?

Fe2O3 + 3 CO → 2 Fe + 3 CO2

7. If 6.57 g of iron are reacted with an excess of hydrochloric acid, then 14.63 g of iron (II) chloride are obtained. What is the percent yield of this reaction? Fe + 2 HCl → FeCl2 + H2

8. A chemist burns 160.0 g of aluminum in excess air. She produces 260.0 g of aluminum oxide. What is the percent yield? 4 Al + 3 O2 → 2 Al2O3

9. If 5.45 g of potassium chlorate are decomposed with a 78.2 % yield, what volume of oxygen was released? 2 KClO3 → 2 KCl + 3 O2

10. What is the actual yield (in grams) of sulfuric acid if 15.0 L of sulfur trioxide reacts with water in a process with a 85.4 % yield? H2O + SO3 → H2SO4

*Problem Set #7: Hydrates*

1. A chemist has a sample of hydrated magnesium sulfite which weighs 1.50 grams. After heating it strongly, he finds that 0.763 grams of water have been driven off. Find the correct formula of the hydrate and name it.

2. A chemist heats a sample of hydrated lithium nitrate which weighs 170 grams and finds, after heating, that it weighs 95.3 grams. Find the correct formula of the hydrate and name it.

3. A chemist heats a sample of baium bromide and finds that it gives off 10.8 grams of water. The compound which is left in the crucible after heating weighs 89.2 grams. What is the formula of the hydrate?

4. What is the molecular formula of dichloroacetic acid if the empirical formula is CHOCl and the molecular mass is 129 g/mole?

5. What is the molecular formula of cyanuric chloride if the empriical formula is CClN and the molecular mass is known to be 184.5 g/mole?

6. The empirical formula for a compound contains one atom of thallium, two atoms of carbon, three atoms of hydrogen and two atoms of oxygen. The molecular mass of the compound is known to be somewhere between 500 and 600 g/mole. What is the molecular formula of the compound?

7. A BINARY hydrocarbon is known to contain 92.3% carbon. If the molecular mass of this compound is known to be between 75 and 80 g/mole, what is the molecular formula?

8. Find the molecular formula of a compound containing three elements if it is known that 26.7% of the compound is phosphorus, 12.1% is nitrogen and the rest is chlorine. The molecular mass is 695 g/mole.

*Unit 8 Review: All of it. Mostly.*

***Choose the best answer to each question.***

\_\_\_ 1. For the reaction Mg + 2 HCl → MgCl2 + H2, calculate the percent yield of magnesium chloride if 100. g of magnesium reacts with an excess of hydrochloric acid to yield 330. g of magnesium chloride.

a. 71.8% b. 74.3% c. 81.6% d. 84.2%

\_\_\_ 2. For the reaction 2 Na + 2 H2O → 2 NaOH + H2, how many grams of hydrogen are produced if 120. g of sodium and 80.0 g of water are available?

a. 4.48 g b. 44.8 g c. 80.0 g d. 200. g

\_\_\_ 3. For the reaction 2Zn + O2 → 2 ZnO, how many grams of zinc oxide can be produced from 100. g each of zinc and oxygen?

a. 100.g b. 124g c. 189g d. 200.g

\_\_\_ 4. A chemist expects to collect 25.0 g of a product based on her calculations. When she performs the experiment, she actually collects 17.3 g of the product. What was her percent yield?

a. 145% b. 59.1% c. 40.9% d. 69.2%

\_\_\_ 5. For the reaction 3 Ca(OH) 2 + 2 H3PO4 → Ca3(PO4)2 + 6 H2O, what mass of calcium phosphate will be formed when 50.0 g of calcium hydroxide is allowed to react with excess phosphoric acid?

a. 3.98g b. 35.8 g c. 69.8 g d. 209 g

\_\_\_ 6. For the reaction 4 Fe + 3O2 → 2 Fe2 O3, what volume of oxygen at STP will be required to form 175 g of iron (III) oxide?

a. 16.3 L b. 24.5 L c. 36.8 L d. 52.3 L

\_\_\_ 7. A reaction is known to have a percent yield of 65.3%. If stoichiometric calculations indicate that a chemist should receive 45.0g of a product, how much will the reaction actually produce?

a. 29.4 g b. 40.8 g c. 59.2 g d. 68.9 g

\_\_\_ 8. In the equation 2 NaCl + H2SO4 → 2 HCl + Na2SO4, the mass of sodium chloride that reacts with 300.0 g of sulfuric acid is

a. 178.7 g b. 300.0 g c. 357.5 g d. 600.0 g

\_\_\_ 9. What volume of N2O can be made from 28.0 L of nitrogen and 13.5 L of oxygen.

The reaction is 2 N2 + O2 → 2 N2O. Assume all gases measured at STP.

a. 13.5 L b. 28.0 L c. 27.0 L d. 6.75 L

\_\_\_ 10. For the reaction C3H8 + 5O2 → 3CO2 + 4 H2O, how many grams of carbon dioxide are produced from the combustion of 250. g of propane?

a. 249 g b. 350.g c. 549 g d. 748 g

\_\_\_ 11. The efficiency of a reaction is measured by the

a. actual yield b. theoretical yield c. percent yield d. limiting reactant

\_\_\_ 12. The calculated amount of product in a reaction is called the

a. actual yield b. theoretical yield c. percent yield d. limiting reactant

\_\_\_ 13. The amount of product that is actually produced in an experiment is called the

a. actual yield b. theoretical yield c. percent yield d. limiting reactant

\_\_\_ 14. The ratio of the actual yield to the theoretical yield is used to calculate the

a. percent yield b. limiting reactant c. products d. excess reactant

16. Ammonia, NH3, is made in the direct reaction between nitrogen and hydrogen. If 280. grams of nitrogen are used, how many grams of hydrogen are also used?

17. Silver bromide, used in photographic film, can be made in the double replacement reaction between silver nitrate and sodium bromide. If you have 2.60 g of sodium bromide, how many grams of silver nitrate are required?

18. Laughing gas, N2O, is made by the decomposition of ammonium nitrate: NH4N O3 → N2O + 2 H2O. If you begin with 155 g of ammonium nitrate, how many liters of laughing gas can you make?

19. Hydrofluoric acid is never sold in glass bottles because glass is composed of silicon dioxide which will react with hydrofluoric acid: SiO2 + 4 HF → SiF4 + 2 H2O. If 375 g of hydrofluoric acid was placed into a glass bottle, how many grams of water will be created?

20. Dinitrogen tetrafluoride can be produced by the reaction of ammonia (NH3) with fluorine gas. If 12.3 liters of ammonia and 26.5 liter of fluorine are allowed to react, how many liters of dinitrogen tetrafluoride will be produced? Also, calculate the mass of excess reactant. 2 NH3 + 5 F2 → N2F4 + 6 HF

21. Methyl alcohol, CH3OH, can be produced in a synthesis reaction between carbon monoxide and hydrogen. If you start with 12.0 grams of hydrogen and 74.5 g of carbon monoxide, how many grams of methyl alcohol are produced? Also, calculate the mass of excess reactant.

22. Disulfur dichloride is a golden yellow liquid with a revolting smell. If is used industrially in the vulcanization of rubber and is prepared by the following reaction: 3 SCl2 + 4 NaF → SF4 + S2Cl2 + 4 NaCl. If you begin with 5.23 g of sulfur dichloride and 6.54 grams of sodium fluoride and isolated 1.19 g of disulfur dichloride, what is your percent yield? Also, calculate the mass of excess reactant.

23. Zinc will react directly with chorine in a synthesis reaction. If you begin with 13.5 g of zinc and the reaction has a 64.3% yield, what mass of zinc chloride was actually produced?

