Experiment 3

Stoichiometry: Following Zinc through Chemical Reactions

Introduction

Combining materials that will cause a chemical reaction is an everyday occurrence. You heat up measured amounts of ingredients when baking a cake, mix the proper amounts of oxygen and gas to make an engine run smoothly, and blow on a flame to increase the intensity of the heat given off in a campfire. All of these examples involve reacting materials to produce a product such as a cake, a smooth-running engine, or a crackling fire. The amount of each component: ingredients in a cake, gas as fuel, or even the oxygen in your breath, will determine exactly how much of the product will be produced. The proportions of the ingredients must be maintained to produce a consistent product. In chemistry we call the process of producing new substances a "**chemical reaction"** and we use a shorthand way of describing a chemical reaction called a **chemical equation.** Chemical equations are useful for knowing exactly which atoms react, what products they will produce, and the relative amounts of each chemical needed to form a particular product. A chemical equation used in chemistry is analogous to following a recipe when cooking.

Mass / Mole Conversions

Unlike in a traditional recipe where we often use standard measurements (teaspoons, cups, etc.), in chemistry, we usually weigh out our "ingredients" by mass. Since some molecules are heavier than others, there will not be a 1:1 ratio between the number molecules of a compound and the relative masses of the compounds reacting, so we can't directly relate the mass of each ingredient to the number of molecules that react. In order to convert the mass of each reacting compound to the number of molecules, we need a conversion factor that takes this difference in weight of the molecules into account. Since molecules are so tiny, we also need to group them together into larger amounts to make the math conversions easier to perform. The unit of **a mole** was developed to allow us to directly convert from a counted number of molecules into groupings with masses large enough to be weighed on a balance; it is similar to having a carton of 12 eggs equaling 1 dozen eggs. **Avogadro's number is a conversion factor where 6.022×10^{23} particles (atoms, molecules, ions, etc.) is set equal to exactly 1 mole of particles**, just like 1 dozen of any item is equal to 12 items. Avogadro's number is used to relate the mass of a compound to the number of moles and gives us a second conversion factor called the **molar mass. The molar mass of a substance is defined as the number of grams of a substance contained in 1 mole (6.022×10^{23} molecules, atoms, ions, etc.) of that substance**.

Calculating Molar Mass

You calculate the molar mass of a compound by adding up the molar masses of all the elements present in your compound. The **molar mass** of an element is listed under its symbol in the periodic table.

Example 1: Calculating molar mass

A molecule of ammonia, NH_3, consists of 1 N atom and 3 H atoms. The molar mass of N is 14.01g/mol and the molar mass of H is 1.008g/mol.

For the calculation of molar mass of ammonia:

$$\frac{14.01g\ N}{1\ mol\ N} + \left(3\ x\ \frac{1.008gH}{1\ mol\ H}\right) = \frac{17.034g\ NH_3}{1\ mole\ NH_3}$$

Example 2: Converting between grams and moles

A molecule of ammonia, NH_3, has a molar mass of 17.034g/mol. The molar mass can be used to convert from mass to moles and from moles to mass.

To go from grams to moles, divide by molar mass:

$$12.00\cancel{g\ NH_3}\left(\frac{1\ \text{mole}\ NH_3}{17.034\cancel{g\ NH_3}}\right) = 0.7045\ \text{mol}\ NH_3$$

To go from moles to grams, multiply by molar mass:

$$1.25\ \cancel{\text{mol}\ NH_3}\left(\frac{17.034g\ NH_3}{1\ \cancel{\text{mole}\ NH_3}}\right) = 21.3g\ NH_3$$

Chemical Equations

Now that you can convert a mass of a material to moles, you need to show the proportions in which molecules will recombine to form products during a chemical reaction. A **chemical equation** is a representation of a reaction, which shows the number of molecules, or moles, of each reactant, followed by an arrow pointing to the products.

$$4NH_3(g) + 5O_2(g) \rightarrow 4NO(g) + 6H_2O(l)$$

For the reaction above, the chemical equation as written could be restated in words as follows: Four moles of gaseous ammonia and five moles of oxygen gas react to form four moles of gaseous nitrogen monoxide and six moles of liquid water.

It is important that the number and type of each atom on the reactants side is exactly equal to the number and type of atoms on the product side. When this is true, we say the chemical equation is "**balanced**". Balancing an equation allows you to account for all atoms on the reactant side being converted to products. There is no net loss of or gain of elements. In the example above, you have 4 N atoms, 12 H atoms and 10 O atoms on the reactant side. The products must also have 4 N atoms, 12 H atoms and 10 O atoms. **All of the reacting atoms must be accounted for in one of the products.** Matter cannot be created or destroyed in chemical reactions.

$$4NH_3(g) + 5O_2(g) \rightarrow 4NO(g) + 6H_2O(l)$$
$$(4N + 12H) + 10O \rightarrow (4N+4O) + (12H+6O)$$
$$4N + 12H + 10O \rightarrow 4N + 12H + 10O$$

Once an equation is balanced, you can predict the amount of product created from a known amount of a reactant since t**he mole ratio between all the materials in the chemical equation is fixed.** Think of a chemical reaction as a recipe in terms of baking. First, if you are making a cake, you must have the correct ratio of all of your ingredients or you will not have a great-looking or tasting cake. Second, if you want to make two cakes, you don't need to do the work twice, you can double ALL of the ingredients (not just the flour!) and bake both cakes at once.

If you double the ingredients (2 x reactants), then you double the cakes (2 x products)

Example 3: Determining moles of a product based on a given number of moles of a reactant.

For the previous example, 4 moles of ammonia, NH_3, always reacts with 5 moles of oxygen, O_2, to produce 4 moles of nitrogen monoxide, NO, and 6 moles of water, H_2O. If you double the amount of each reactant, (8 moles of ammonia, and 10 moles of oxygen), you will double the amount of product produced. (8 moles nitrogen monoxide and 12 moles water).

$$2 \times (4NH_3(g) + 5O_2(g) \rightarrow 4NO(g) + 6H_2O(l))$$
$$8NH_3(g) + 10O_2(g) \rightarrow 8NO(g) + 12H_2O(l)$$

Now you have the means to predict the amount of a product produced from any given amount of a reactant using the mole ratio in the chemical reaction.

<u>Dimensional Analysis</u>

A single conversion between grams and moles or predicting the moles of product from a given number of moles of a reactant is relatively simple. However, many of the calculations you will need to perform in this course will involve many conversions to get from the initial measurement of a reactant to the final amount of a compound produced.

Dimensional analysis (see notes on experiment 1) is a method of keeping track of your units when making a series of unit conversions. It allows you to check your units and to be sure your calculations will be correct BEFORE you actually do any math. Use the following rules and example and you should be able to convert from 1 unit to another without errors.

<u>Example 4: Calculating the mass of H_2O produced from 1.5g NH_3</u>

1. Always label every value you have with both the unit and chemical associated with it.

   1.5 g NH_3

   value unit chemical

2. Write down all given values and conversion factors that you think you may need.

   a. 1.5g NH_3
   b. 17.034g NH_3/1 mol NH_3
   c. 30.036g NO/1 mol NO

3. Write down the balanced chemical equation you will need to go from one chemical to another.

   $$4NH_3(g) + 5O_2(g) \rightarrow 4NO(g) + 6H_2O(l)$$

4. Write down the unit and chemical you want to end up with. (g H_2O)

5. Set up a series of steps so that an added unit cancels an unwanted unit in the previous step.
   (a) Use a horizontal line to divide numerator and denominator, not a slash.
   (b) Use a 1 in the numerator or denominator as needed to keep track of top and bottom units.
   (c) Invert the conversion factor when needed to cancel top and bottom units.
   (d) Keep adding steps until you only have the final unit or units you want in the answer.

6. Cancel units top and bottom and check that you end up with the same unit and chemical that you originally wrote down on the left side of the equation. Then do out the math.

$$\boxed{g\,H_2O} = \frac{1.5g\cancel{NH_3}}{1} x \frac{1\cancel{mol\ NH_3}}{17.034\cancel{gNH_3}} x \frac{6\ \cancel{mol\ H_2O}}{4\cancel{mol\ NH_3}} x \frac{18.016gH_2O}{1\cancel{molH_2O}} = \frac{2.4\ \boxed{gH_2O}}{1}$$

Limiting Reagent

Knowing the fixed ratio of the chemicals in a chemical equation allows you to predict the amount of product that can be produced from a known amount of each reactant. In the reaction shown below, 4 moles of ammonia, NH_3, produce 4 moles of nitrogen monoxide, NO, as long as at least 5 moles of oxygen are present.

$$\underline{4}NH_3(g) + 5O_2(g) \rightarrow \underline{4}NO(g) + 6H_2O(l)$$

This brings us to the concept of a **limiting reagent**. A limiting reagent is the reactant that will run out first. So, if you have 10 moles of O_2 but only 4 moles of NH_3, you still can only produce 4 moles of NO. All of the ammonia NH_3, will be used up, leaving extra oxygen left over. Therefore, the ammonia, NH_3, is limiting how much product that can be produced. Thus, ammonia is the **limiting reagent**, in this reaction.

$$4NH_3(g) + \underline{10}O_2(g) \rightarrow 4NO(g) + 6H_2O(l) + \underline{5}O_2(g)$$

Theoretical, Actual, & Percent Yield

Theoretically, we can predict exactly how much material can be produced in an equation by looking at the limiting reagent. We call this the **theoretical yield**. In theory, we could collect 4 moles of NO gas if 4 moles of NH_3 are used as a reactant and we have extra O_2 (g).

$$4NH_3(g) + \underline{10}O_2(g) \rightarrow 4NO(g) + 6H_2O(l) + \underline{5}O_2(g)$$

However, in real life, this is not usually the amount of material that is recovered. Frequently some product is lost or the limiting reagent is not fully reacted, resulting in a measured mass of product being less than what was predicted mathematically. This experimentally determined mass of the product is the **actual yield**. The actual yield is the amount of product actually measured at the end of a reaction.

Finally, we want to know just how good our experimental results really are. Did we generate a lot of product based on the amount of the limiting reagent used, or was our actual yield pretty low, indicating that we lost product somewhere during the course of the reaction? The measure of how close you were to collecting 100% of the product is called the **percent yield**. It is the ratio of the actual yield to the theoretical yield multiplied by 100 to show it is a percentage.

$$\text{Percent Yield} = \left(\frac{\text{Actual Yield}}{\text{Theoretical Yield}}\right) \text{x } 100$$

In the above reaction, we should theoretically be able to produce 2 moles of NO if 2 moles of NH_3 are used. This corresponds to 60 grams of NO. However, due to a leak in the container used to collect the NO, we only collected 45 grams of NO when the experiment was finished. Our percent yield is:

Theoretical yield of NO:	60g NO
Actual yield of NO:	45g NO
Percent yield:	(45g NO / 60g NO) x 100 = 75%

The percent yield in this example shows us that we collected 75% of the NO possible, but lost 25% of the possible product during the course of the reaction. Thus, a percent yield is an indication of the success of the reaction.

In Your Experiment

You will be decomposing zinc carbonate through a series of reactions to eventually form zinc oxide.

The reactions you will do are listed below:

1. Add acid, the $ZnCO_3(s)$ dissolves and you see gas bubbles from the carbon dioxide formed.

$$ZnCO_3(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + CO_2(g) + H_2O(l)$$

2. Add sodium hydroxide, NaOH, and you see a gelatinous precipitate, $Zn(OH)_2$.

$$ZnCl_2(aq) + 2NaOH(aq) \rightarrow Zn(OH)_2(s) + 2NaCl(aq)$$

3. Heat the zinc hydroxide to remove water. Note the steam (H_2O) coming off as solid white ZnO is formed.

$$Zn(OH)_2(s) + heat \rightarrow ZnO(s) + H_2O(g)$$

Thus, the overall reaction for the decomposition of zinc carbonate is as follows:

$$\mathbf{ZnCO_3(s) \rightarrow ZnO(s) + CO_2(g)}$$

Your goal in this experiment is to recover as much ZnO as possible. To evaluate your level of success, you must calculate the theoretical yield of ZnO from the original amount of $ZnCO_3$ so that you can then calculate your percent yield. The steps are as follows:

1. Weigh out a mass of $ZnCO_3$
2. Convert the mass of $ZnCO_3$ to moles of $ZnCO_3$ using the molar mass of $ZnCO_3$
3. Convert the moles of $ZnCO_3$ to moles ZnO using the mole ratio in the overall reaction. (1:1)
4. Convert the moles of ZnO to grams of ZnO using the molar mass of ZnO.
5. This mass of ZnO found using the original mass of $ZnCO_3$ is the theoretical yield of ZnO when determining the percent yield of ZnO.

Your % yield is primarily based on how well you perform the reaction, so work carefully! This experiment should result in percent yields well over 80% and we often see yields of nearly 100%.

Lab Precautions

<u>Chemical Hazards</u>

<u>Zinc Carbonate</u>

NFPA RATING: HEALTH: 0 FLAMMABILITY: 0 REACTIVITY: 0
EYE EXPOSURE: Flush with copious amounts of water for at least 15 minutes. Assure adequate flushing by separating the eyelids with fingers.

<u>3M Hydrochloric Acid Solution</u>

NFPA RATING: HEALTH: 2 FLAMMABILITY: 0 REACTIVITY: 0
DERMAL EXPOSURE:
Caustic. WEAR GLOVES to prevent exposure. In case of contact, wash off with plenty of water.
EYE EXPOSURE:
Caustic. In case of contact with eyes, flush with copious amounts of water for at least 15 minutes. Assure adequate flushing by separating the eyelids with fingers. Contact your TA immediately for medical help.
REACTIVITY
Will react with strong bases, work with caution.

<u>3M Sodium Hydroxide Solution</u>

NFPA RATING: HEALTH: 2 FLAMMABILITY: 0 REACTIVITY: 1
ORAL EXPOSURE
Caustic. If swallowed, wash out mouth with water provided person is conscious. Do not induce vomiting. Contact your TA immediately.
DERMAL EXPOSURE
Caustic. WEAR GLOVES. In case of extensive skin contact, flush with copious amounts of water for at least 15 minutes. Remove contaminated clothing and shoes. Contact your TA immediately.
EYE EXPOSURE
Caustic. In case of contact with eyes, flush with copious amounts of water for at least 15 minutes. Assure adequate flushing by separating the eyelids with fingers. Contact your TA immediately for medical help.
REACTIVITY
Will react with strong acids, work with caution.

<u>Chemical Disposal</u>

Dispose of all chemicals in the liquid waste container in the lab. Solid ZnO will be converted to aqueous $ZnCl_2$ before disposal.

Equipment Procedures

Using a Hot Plate/Stirrer

The hot plates used in lab have a white ceramic top that will heat up very quickly. Unlike a stove burner, these hot plates will not get red as they heat up, thus a heated hot plate looks exactly like a cold hot plate. Please use the following precautions when using a hot plate in lab.

1. Make sure the power cord is not touching the hot plate at any time during use.
2. Keep the hot plate on the lowest setting possible for your experiment.
3. Do not leave your hotplate unattended.
4. Clear glassware should only be used on a hot plate if it contains a liquid or you are carefully monitoring a drying solid.
5. When drying a solid, the hot plate should be on a **LOW** setting.
6. Clean up any chemicals that may spill on the hot plate.
7. The hot plates also have a stir setting. Make sure you are using the correct dial before notifying your TA that your hot plate does not work.
8. Turn off and unplug the hot plate before leaving the lab.

There are 2 dials on the hot plate. One is for heating and the other for stirring. The exact temperature is not given, so use care when moving the dial to a higher setting. The hot plates do take some time to warm up and moving the setting to high immediately, will NOT make the hot plate warm up more quickly. Use a setting of 2-3 when drying a solid, especially in a beaker. Any higher than that and the chemicals may pop out of the beaker or you may crack the bottom of your beaker (and need to pay for a new one).

Using a Centrifuge

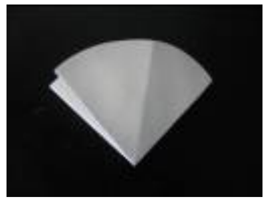
A **centrifuge** is a piece of equipment usually used to separate a heavier solid (the **precipitate**) from a liquid in a suspension. A **suspension** is heterogeneous mixture where a solid is held within a liquid, but is not dissolved. A centrifuge uses centrifugal force to drive the heavier components of a mixture to the bottom of the tube, leaving the less dense components on top of the solid. Follow the instructions below carefully to avoid damage to the centrifuge or loss of your product.

1. Pour your solution into either a plastic centrifuge tube or a Kimax or Pyrex test tube. Do not use a glass culture tube without markings, they tend to break. You are responsible for replacing any broken test tubes, so be sure to check it before you put it in the centrifuge.
2. Fill a second tube to the same volume as the test tube containing your solution.
3. Open the top of the centrifuge.
4. Insert the two test tubes directly across from each other so they will balance out.
5. Close the lid and turn on the centrifuge to the setting given in your lab.
6. Wait for the centrifuge to come to a full stop.
7. Remove your test tubes.
8. Remove the "**supernatant**" (liquid on top of **precipitate**) with a plastic pipet and transfer to a separate container.
9. Your solid and liquid are now separated.
10. Avoid bumping the test tubes when they are removed to minimize remixing of the supernatant and the precipitate.

Filtering a Solution with a Funnel and Filter Paper (Gravity Filtration)

Filtration is a mechanical means of separating a solid and a liquid. There are many types of filtration; these instructions are for gravity filtration only.

1. Insert a funnel into a beaker or Erlenmeyer flask so that the tip of the funnel does not touch the bottom of the glassware.
2. Fold a round filter paper in half, then in half again to create a quarter sheet size of filter paper.



3. Separate 1 fold of the filter paper to create a cone shape.
4. Insert the cone shape into the funnel.

5. Mix your solid/liquid suspension thoroughly and pour into the filter paper cone. Do not allow liquid to rise over the top of the filter paper.
6. Keep pouring your suspension into the cone until it is gone.
7. Use a wash bottle of distilled water to rinse out any solid that may not have been transferred to the funnel. Use as little water as possible to remove solid. Every drop of water you use has to filter through the paper adding to your lab time.
8. Allow the solution to drip through the filter paper until no more drops are seen.
9. At this time, you can scrape the precipitate off the filter paper into another container, usually a beaker or watch glass to be dried. Be sure to scrape off as much residue as possible. If necessary, a small amount of water can be used to get the residue into the collection container. However, be aware the more water you use to remove the solid will result in a longer drying time.

Experiment 3: Procedures and Data Sheet

(Submit as part of your informal report)

Name: ____ Date: _______ Section: _______

TA Signature: ___________________________________

All data must be written in pen at the time it is collected. **Pencil is not allowed!!**
Record all measurements with the correct number of significant figures and units.
TA signature & TA initials on any changes made to the data are required or the data is invalid.

Mass of $ZnCO_3$

1. Place a clean, dry, 4" test tube in a 50mL beaker on the balance.
2. Record the mass of the test tube and beaker.
3. Weigh out between 0.200g and 0.250g of $ZnCO_3$ into the test tube. Be careful, the chemical is very fluffy. Clean up any chemical mess! You will be graded on neatness in the lab.
4. Record the exact mass of the $ZnCO_3$, test tube and beaker.

Mass of test tube and beaker ___________

Mass of test tube beaker and $ZnCO_3$ ___________

Production of $ZnCl_2$

1. Slowly add 2mL (40 drops) of 3M HCl into the test tube containing the $ZnCO_3$. You will see bubbles of CO_2 coming out of the solution. If necessary, stir the mixture with a spatula until all traces of $ZnCO_3$ have disappeared. Your insoluble zinc carbonate has now been converted to soluble zinc chloride, $ZnCl_2$.
   a. **Caution: HCl and NaOH are both caustic chemicals so avoid skin and eye contact. If there is any skin contact wash immediately with lots of water and notify your TA. Always wear eye protection.**
2. Record all observations of this reaction.

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Production of $Zn(OH)_2$

1. Add approximately 2.5 mL (50 drops) of 3M NaOH to the test tube to convert the zinc chloride to zinc hydroxide, $Zn(OH)_2$. Stir the mixture with a spatula to make sure all of the zinc reacts. You will now have a test tube with a white gelatinous precipitate. This is the $Zn(OH)_2$. Scrape as much of the precipitate off the spatula as possible back into the test tube. Set the test tube upright in a beaker to settle.
2. Record all observations of this reaction.

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Production of ZnO

1. Fill a second 4" test tube with a volume of water equal to the volume of the sample mixture in your test tube. Place the test tube containing your sample in an empty hole in the centrifuge. Set the test tube containing the water in the hole in the centrifuge that is directly across from the test tube containing your mixture. This second test tube will act as a balance in the centrifuge and prevent the centrifuge from moving sideways on the lab bench once the rotor begins to spin. Close the cover and set the dial to centrifuge your sample for 2 minutes.
2. Carefully remove both test tubes from the centrifuge.
3. Remove the supernatant solution from your sample with a plastic pipet. Discard this solution in a 250ml beaker as waste. Leave about ½ mL of the solution in the test tube so that you do not disturb the precipitate.
4. Add ~2 mL of water with your eyedropper to the test tube. Try to wash all of the precipitate that has accumulated on the side of the test tube into the solution. Use your spatula to stir the precipitate so that all of the precipitate can be washed. This will remove all traces of NaOH and HCl that might still be in the solution. Wipe as much of the precipitate off the spatula as possible into the test tube.
5. Set up your filtration apparatus.
6. Stir up the precipitate and quickly pour the $Zn(OH)_2$ suspension into your filter paper in your funnel. Be sure not to go over the top of the filter paper. Add ~1mL of water to the empty test tube to remove any traces of the $Zn(OH)_2$. Pour this wash into the funnel as well.
7. Allow your solution to filter until no liquid is left in the funnel.
8. Place a 50 mL beaker on the balance and record the mass of the beaker.

   Mass of 50 mL beaker ____________

9. Remove the filter paper from the funnel and gently scrape the solid $Zn(OH)_2$ into the pre-weighed beaker. Use a **<u>very</u>** small amount of water from your squirt bottle to remove as much $Zn(OH)_2$ as possible from the filter paper.
10. Place the beaker on a hot plate on **low heat** and evaporate the water. As the solution evaporates, the $Zn(OH)_2$ loses water and is converted to ZnO.
11. Record all observations of this reaction

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12. Wait 10-15 minutes and remove the beaker from the hotplate with a folded paper towel to protect you from the heat. Your ZnO should be a dry, cracked solid on the bottom of the beaker. If it is still damp, keep heating until it appears dry.
13. Cool 5 minutes.
14. Scrape the ZnO from the sides and bottom of the beaker with a spatula. Use your spatula to crush the ZnO into a powder. You should now have a dry beaker with a loose, powdery looking solid on the bottom.
15. Weigh the beaker containing the ZnO on the same balance that you used for the initial weighing of the beaker. Record the weight of beaker and ZnO.
16. Return the beaker to the hot plate and heat for 5 more minutes. Remove from heat, cool and reweigh. If the weight is within 0.020g of the original mass recorded from the previous step, record this mass as the final mass of beaker and ZnO. If it is not within 0.020g, return the beaker to the hot plate for another 5-10 minutes. Repeat reheating, cooling and weighing until you reach a constant mass or the lab period is over, whichever comes first.

Mass of 50 mL beaker and ZnO ____________

Mass of 50 mL beaker and ZnO ____________

Mass of 50 mL beaker and ZnO ____________

Cleanup

1. Once you have recorded the final mass of the beaker and the ZnO, add ~ 1 mL of 3M HCl to the beaker to dissolve the solid and convert back to $ZnCl_2$.
2. Swirl the contents until NO traces of solid ZnO remain. Add more HCl if needed.
3. Pour the zinc chloride ($ZnCl_2$) solution into the waste container.
4. Wash glassware and your spatula with water to remove all chemical residue. Use a few drops of HCl to remove stubborn residue. See your TA if you cannot get your equipment completely clean. **If anything is left dirty, it will be reported as unusable by the next student and you will be charged for it.**
5. Wipe down balance and lab area with a damp paper towel to be sure no chemical residue remains at your station.

Experiment 3: Data Rubric (20pts)

Points

Data are neat and legible	5pts	__________pts
Significant figures (>80% correct)	3pts	__________pts
Units (>80% correct)	2pts	__________pts
All data are present and make sense	10pts	__________pts

Deductions (sliding scale based on TA discretion)

Lab area left unclean	-20pts	__________pts
Improper waste disposal	-20pts	__________pts
Disruptive behavior	-20pts	__________pts
Lab coat or safety glasses removed while in lab	-20pts	__________pts
Data sheet is missing TA signature	-20pts	__________pts
Other: ____		__________pts

Comments: ____

Grade for Data Sheet __________**pts**

Experiment 3: Results Table

(Submit as part of your informal report)

Name: ______________________________________ Date: ________ Section: ________

Record all results with the correct number of significant figures and units.

Mass $ZnCO_3$	
Molar Mass of $ZnCO_3$	
Moles $ZnCO_3$	
Theoretical moles ZnO	
Molar Mass of ZnO	
Theoretical mass of ZnO	
Actual mass of ZnO collected	
Percent yield of ZnO	

Experiment 3: Results Table Rubric (20pts)

Points

Tables are neat and legible	5pts	___________pts
Significant figures (>80% correct)	3pts	___________pts
Units (>80% correct)	2pts	___________pts
All results are present and make sense	10pts	___________pts

Deductions (sliding based on TA discretion)

Results to not match data	-20pts	___________pts
Plagiarism!!! Results are identical to another student	**-100pts**	**___________pts**

Other: ____ ___________pts

Comments: ____

Grade for Results Table **___________pts**

Experiment 3: Calculations

(Submit as part of your informal report)

Mass of $ZnCO_3$

Subtract the mass of the test tube from the mass of the test tube and the $ZnCO_3$ to get the mass of $ZnCO_3$.

Mass of $ZnCO_3$ ____________

Molar Mass of $ZnCO_3$

Use the molar masses in the periodic table to calculate the molar mass of $ZnCO_3$.

Molar mass of $ZnCO_3$ ____________

Moles of $ZnCO_3$

Use the mass and molar mass of $ZnCO_3$ to calculate the moles of $ZnCO_3$. Use dimensional analysis to make sure that your math is correct – grams should cancel leaving moles.

Moles of $ZnCO_3$ ____________

Theoretical Moles of ZnO

Use the moles of $ZnCO_3$ and the mole ratio from the balanced chemical equation:

$$ZnCO_3(s) \rightarrow ZnO(s) + CO_2(g)$$

to calculate the theoretical number of moles of ZnO produced in your reaction.

Moles of ZnO ____________

Molar Mass of ZnO

Use the molar masses in the periodic table to calculate the molar mass of ZnO.

Molar mass of ZnO ___________

Theoretical Mass of ZnO

Use the molar mass of ZnO and the theoretical number of moles of ZnO produced to calculate the theoretical mass of ZnO. Use dimensional analysis to make sure that your math is correct – moles should cancel leaving grams.

Theoretical mass of ZnO ___________

Actual Mass of ZnO

Subtract the mass of the beaker from the mass of the beaker and the ZnO.

Actual mass of ZnO ___________

Percent Yield of ZnO

Use the actual mass of ZnO and the theoretical mass of ZnO to calculate the percent yield.

$$\text{Percent Yield} = \left(\frac{\text{Actual Yield}}{\text{Theoretical Yield}}\right) \times 100$$

Percent yield of ZnO ___________

Experiment 3 Questions:

Submit as part of your informal report

1. Evaluate the success of your experiment based on your percent yield. If your percent yield was less than 100%, describe at least two possible areas where you might have lost material. If your percent was over 100%, explain what you feel might have gone wrong with your experiment.

2. Imagine that you were conducting an experiment in which you were making sodium chloride from copper (II) chloride according to the following reaction:

$$CuCl_2 + 2Na \rightarrow 2NaCl + Cu$$

   If you started with 0.236g of $CuCl_2$, what would be your theoretical yield of NaCl?

   Answer: ___________________

Experiment 3: Prelab Worksheet

(Submit via Brightspace before coming to lab)

Name: ______________________________ Date: _______ Section: _______ Grade: _______

All information needed to complete this worksheet can be found in the pre-lab information and calculations sections of the lab manual. Read this introductory material first!

- Record all values with the <u>correct number of significant figures and units</u>.
- Place all answers on the line when provided.
- Show calculations for any numerical answers; **work must be shown to receive credit**.
- See any 102 TA via webex before your prelab is due if you have any questions.
- Each question is worth 2 points.

Write the balanced chemical equation associated with each of these 3 reactions. Include the phases of all chemicals. (See lab manual)

1. Add aqueous hydrochloric acid to dissolve the zinc carbonate forming gaseous carbon dioxide, an aqueous solution of zinc chloride, and liquid water.

   ____

2. Add aqueous sodium hydroxide to the zinc chloride solution and centrifuge to form gelatinous zinc hydroxide and aqueous sodium chloride.

   ____

3. Heat the zinc hydroxide to remove water leaving solid zinc oxide in the beaker.

   ____

4. Name a “caustic” chemical used in this experiment.

   Caustic chemical ___________

5. Name two techniques that are used to separate the zinc hydroxide from the aqueous salt solution in this experiment.

   ____

   ____

6. What can happen to your beaker if you heat it too quickly on your hot plate?

____

7. Calculate the molar mass of zinc hydroxide. Show all work.

Molar mass of zinc hydroxide_____________

Use the following information to answer questions 8 & 9: You started by weighing out 0.552g of zinc carbonate. After the experiment was finished, the mass of zinc oxide collected was 0.225g.

8. Calculate the theoretical mass of ZnO that could be formed from 0.552g of $ZnCO_3$ in this experiment. Show all work, including how your units cancel.

Theoretical mass of ZnO ____________

9. Calculate the percent yield of ZnO after the reaction. Show all work.

Percent yield of ZnO ____________

10. Will your hot plate turn red when it is hot?

____