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Empirical Formula of Magnesium Oxide

Answer all questions in this text document. Submit your document to be graded. You may print, hand write and scan. Or you can word process directly on this document.

Introduction

Chemical formulas are used to describe the atomic ratio of atoms in a compound. In *molecular formulas,* the ratios relate to the molecule as a whole. Benzene, for example, has a molecular formula of C_6H_6, which means that there are 6 carbon atoms and 6 hydrogen atoms in each benzene molecule. In other words, the ratio of carbon to hydrogen is 6:6.

Empirical formulas give the simplest whole-number ratio of the atoms in a compound. For benzene, this is CH (carbon to hydrogen ratio of 1:1). For propylene, C_3H_6, the empirical formula is CH_2. Empirical formulas are an important part of analytical chemistry because they are simple to determine experimentally (or, *empirically*).

Say you have 13.8 g of a compound that contains only nitrogen and oxygen. By decomposing the compound, you determine that it contains 4.2 g of nitrogen. How would you determine the empirical formula?

Remember that atomic ratios and molar ratios are the same. If you find the molar ratio of the nitrogen and oxygen, it will be the same as the atomic ratio of the nitrogen and oxygen. The atomic ratio is your empirical formula.

So, first, find the number of moles of nitrogen and oxygen in your sample:

$$4.2\text{ g N} \times \frac{1\text{ mol N}}{14.01\text{ g N}} = 0.30\text{ mol N}$$

$$(13.8\text{ g} - 4.2\text{ g})\text{ O} \times \frac{1\text{ mol O}}{16.00\text{ g O}} = 0.60\text{ mol O}$$

Divide each of those by the smallest number to get the formula:

N: 0.30/0.30 = 1

O: 0.60/0.30 = 2

Thus, the empirical formula of your compound is NO_2.

In the following experiment, you will determine the empirical formula of magnesium oxide, using the same thinking as the nitrogen dioxide determination you just read about here.

Pre-Lab Questions

Heating 1.29 g of gallium gives a gallium-oxygen compound with a mass of 1.73 g. Calculate the empirical formula of the compound.

Step 1: Figure out how many grams of oxygen are in the 1.73 grams of gallium-oxygen compound by subtracting.

1.73 − 1.29 = 0.440 grams of oxygen

Step 2: Convert the 1.29 g of gallium to moles:

$$1.29\text{g Ga} \times \frac{1\text{ mol Ga}}{69.72\text{g Ga}} = 0.019\text{ mol Ga}$$

Step 3: Convert the grams of oxygen you calculated in step 1 to moles:

$$0.440\text{g O} \times \frac{1\text{ mol O}}{16.00\text{g O}} = 0.028\text{ mol O}$$

Step 4: Figure out the formula.

$$\frac{0.019}{0.019} = 1 \text{ atom Ga} \qquad \frac{0.028}{0.019} = 1.47 \text{ atoms O}$$

× 2   × 2

2 atoms Gallium   2.97

3 atoms Oxygen

Empirical Formula: Ga_2O_3

Name of Compound: Gallium (III) oxide

Empirical Formula of Magnesium Hydroxide
Procedures and Data

Watch the following video and answer the following questions> I suggest you use a split screen so that you can watch the experiment as you collect the data. https://www.youtube.com/watch?v=OuFqtxZJRvM

STEP 1: Determine the mass of the Magnesium. Record the data in the table below with all Sig. fig. and units:

Mass of Crucible and Lid	31.064g
Mass of Crucible and Lid and Magnesium	31.634g
Mass of Magnesium (show calculation)	0.57g

STEP 2: Heat the Magnesium

What are two safety warnings?

- When lighting a bunsen burner, be sure that all hair and clothing are clear of the flame.
- Do not look directly at the burning magnesium.

The author provides several tips for successfully heating the magnesium. List two here:

- The ribbon should be loosely coiled in the base of the crucible.
- Keep the bunsen burner close to the crucible when heating.

STEP 3: Run the reaction to completion

Write the reaction equation that is occurring while you are heating magnesium with the Bunsen Burner. Remember, oxygen is diatomic. Balance the equation.

- $2Mg + O_2 \rightarrow 2MgO$

How do you know when the reaction is done, and the magnesium is completely changed to magnesium oxide?

- The magnesium will no longer ignite when exposed to oxygen.

NEXT STEPS:

List the next five steps with a brief rationale for each step (why is it done?)

- (After cooling) Drop water into the Magnesium oxide → Water will supply hydrogen
- Stir the water into the paste → this makes a uniform mix.
- Reheat the crucible to boil the water → Agitates the substances to react
- Weigh the crucible with the lid. Can be compared against the previous measurement to
- Clean and scrub the crucible. Ensures that the crucible can be reused.

stir the paste

FINAL RESULTS

Observation of product	Light gray solid paste
Mass of product and crucible	31.97g
Mass of crucible (from first data table)	31.064g
Mass of product (show calculation).	0.906g

DATA ANALYSIS

1) Determine the empirical formula of MgO based on your data recorded. Follow the steps used in the introduction to find the empirical formula of NO_2. (In brief: Subtract Mg mass from Step 1 from product mass of Final results step to calculate mass of oxygen, then convert the grams of Mg to moles, convert the grams of O to moles, divide by the smallest number to get whole numbers, write the formula).

mass of oxygen = 0.906 − 0.570 = 0.336

$0.570 \cdot \frac{1 \text{ mol Mg}}{24.31 \text{ g Mg}} = 0.023$ mol Mg

$0.336 \cdot \frac{1 \text{ mol Mg}}{16.00 \text{ g Mg}} = 0.021$ mol O

$\frac{0.023}{0.021} = 1.095$ atoms

$\frac{0.021}{0.021} = 1$ atoms

$Mg_{219}O_{200}$?

This is roughly 1:1, which would be MgO

2) Determine the theoretical formula of MgO using the periodic table. To do this: find the most stable ion of magnesium and the most stable ion of oxygen. Use that to predict the formula for magnesium oxide:
   - Mg ion → Mg^{2+}
   - O ion → O_2^{2-}
   - Formula for Magnesium oxide MgO

3) Compare what you got in #1 to what you should have got in #2. How close were you? How might you explain the difference?

The ratio should be 1:1, but perhaps due to a rounding error or some imprecision in the measurement, the calculated mass of the Magnesium was not equal to that of oxygen. Alternatively, the presence of Hydrogen in the afterproduct could also affect the data.

Post-Lab Question

1) **Hypothetical Data for Analysis**

An oxide of iron is 69.94% iron by mass. Calculate its empirical formula and determine the name of the ionic compound. (Hint: assume you have a 100 g sample of the oxide of iron. How many grams would be iron? Then how many grams would be oxygen? Convert those to moles. Find the ratio. Find the nearest whole number ratio). Then determine its name.

(1) 69.94 grams Fe
30.06 grams Oxygen

$69.94 \cdot \frac{1}{55.85} = 1.252$ mol Fe

$30.06 \cdot \frac{1}{16} = 1.879$ mol O

$\frac{1.879}{1.252} = 1.5$

Fe : O
1 : 1.5
2 : 3

Fe_2O_3

Empirical Formula: Fe_2O_3

Name of Compound: Iron (III) oxide

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Module 6: Percent Yield of NaCl

Introduction

Reaction stoichiometry is used to relate the amounts of reactants and products in a chemical reaction. For example, consider the combustion of propane:

$$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$$

This equation tells you that for every 1 mol C_3H_8 combusted, 5 mol O_2 must react and 3 mol of CO_2 and 4 mol H_2O must be produced. No matter how much propane is consumed, these ratios must always be the same. Thus, it is possible to predict the number of moles of CO_2 that would be produced if 40.0 mol O_2 is reacted:

$$40.0 \text{ mol } O_2 * \frac{3 \text{ mol } CO_2}{5 \text{ mol } O_2} = 24.0 \text{ mol } CO_2$$

It is also possible to predict the mass of CO_2 that would be produced from the combustion of 61.0 g of propane:

$$61.0 \text{ g } C_3H_8 * \frac{1 \text{ mol } C_3H_8 \quad 3 \text{ mol } CO_2 \quad 44.01 \text{ g } CO_2}{44.11 \text{ g } C_3H_8 \quad 1 \text{ mol } C_3H_8 \quad 1 \text{ mol } CO_2} = 183 \text{ g } CO_2$$

183 g is the theoretical yield of CO_2. The amount that SHOULD be made according to theory. The actual amount of CO_2 collected will almost always be less. There are a variety of reasons why the actual yield is less than the percent yield. Sometimes the reaction is not allowed to finish. Sometimes a different reaction occurs, causing byproducts. Sometimes you cannot collect all of the product generated. To account for these difficulties, it is common to report the percent yield:

$$\text{percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} * 100\%$$

If 125 g of CO_2 were actually collected from the above reaction, the percent yield would be:

$$68.3\% = \frac{125 \text{ g } CO_2}{183 \text{ g } CO_2} * 100\%$$

Pre-Lab Questions

You will be observing a reaction in this lab. As with most reactions, this reaction occurs by the breaking of some bonds in the reactants and the formation of new bonds in the products. During the reaction, the gas that is produced will fizz away. After the reaction is complete, one of the products can be removed. Then you will determine the mass of the other product – this will be your actual (or experimental) yield. You will compare this value to what you should have got (theoretical yield) to calculate your % yield.

1. Balance the following equation: $\mathbf{NaHCO_3 + HCl \rightarrow NaCl + H_2O + CO_2}$

	Na	H	C	O	Cl
R	1	2	2	3	1
P	1	2	2	3	1

Already Balanced

2. List the reactants that will be put into the beaker:
   - Baking Soda (sodium bicarbonate or sodium hydrogen carbonate)
   - Hydrochloric acid

3. List the products that will be formed in this reaction:
   - Table salt (sodium chloride)
   - Water (dihydrogen monoxide)
   - Carbon dioxide

4. Which product will fizz away?

   Carbon Dioxide

5. How can you get rid of another one of the products so that there is only one left?

   Boil away the water to leave the salt

6. How many grams of NaCl should be produced if you start with **1.0 g** of sodium bicarbonate. This conversion will require three conversion factors. *(Ref: Example 8.3 and 8.6, chapter 8 of textbook)*.

Molar Mass of $NaHCO_3 = 22.99 + 1.01 + 12.01 + (16.00)3 = 84.01g$

Molar Mass of $NaCl = 22.99 + 35.45 = 58.44$

$$1.0g\ NaHCO_3 \cdot \frac{1\ mol\ NaHCO_3}{84.01g\ NaHCO_3} \cdot \frac{1\ mol\ NaCl}{1\ mol\ NaHCO_3} \cdot \frac{58.44g\ NaCl}{1\ mol\ NaCl} = 0.696g\ NaCl$$

7. How many molecules of CO2 should be produced if you start with **1.0 g** of sodium bicarbonate. This conversion will require three conversion factors. *(Ref: Example 8.3, chapter 8 and Example 6.1, chapter 6 of textbook)*.

molar mass of $NaHCO_3 = 84.01g$

molar mass of $CO_2 = 12.01 + (16)2 = 44.01$

$$1.0g\ NaHCO_3 \cdot \frac{1\ mol\ NaHCO_3}{84.01g\ NaHCO_3} \cdot \frac{1\ mol\ CO_2}{1\ mol\ NaHCO_3} \cdot \frac{6.022\times10^{23}}{1\ mol\ CO_2} = 7.168\times10^{21}$$

% Yield of NaCl

Procedures and Data

Watch the following video and answer the following questions. I suggest you use a split screen so that you can watch the experiment as you collect the data.
https://www.youtube.com/watch?v=vjVrIFScsls

STEP 1: Determine the mass of $NaHCO_3$.
Record what the data should be for the second measurement in the table below and then do the subtraction to show that you have 2.00 g of $NaHCO_3$ in the beaker.

Mass of Beaker	**145.55 g**
Mass of Beaker and $NaHCO_3$	147.55g
Mass of $NaHCO_3$ (show calculation)	2.00g

STEP 2: Add chemicals.
Write down the next two reagents added and the rationale (reason) for them.
- Bromothymol Blue – used to indicate the acidity of the substance
- Water – used to help dissolve the $NaHCO_3$

Notice that the color of the solution is blue.
What does this mean? The solution is basic
What would it mean if it were yellow? The solution is acidic

STEP 3: Run the reaction to completion
Record your observations in the table below. Explain what the observations indicate (inferences).

Observations upon addition of first four drops of HCl	Inferences
Bubbles begin to form	CO_2 gas is being released
The color turns yellow	The solution is now acidic

How will you know when the reaction is complete (three indications).
- The solution turns yellow
- All of the $NaHCO_3$ is dissolved
- No more CO_2 is being released (no more bubbles)

FINAL STEPS: List the last two steps needed to isolate NaCl.
- Rinse the beaker to dissolve any remaining $NaHCO_3$
- Boil away any excess water

FINAL PRODUCT: The video does not show the result after evaporation. If it had, you would see salt (NaCl) alone in the beaker. The following data table provides the data you need to calculate the amount of salt you produced:

Mass of beaker alone (see data table above)	145.55g
Mass of NaCl and beaker	**146.83 g**
Mass of NaCl (show subtraction)	1.28g

% Yield of NaCl

Data Analysis

1. What was the mass of $NaHCO_3$ you started with (see first data table)? 2.00 g

2. The mass of NaCl that should be produced is called the Theoretical yield because it is the amount that should be produced according to theory. What is the mass of NaCl that should be produced in theory? (Show your calculation. Start with the mass of $NaHCO_3$ from above and use conversion factors like you did in the pre-lab worksheet).

$$2.00 g\ NaHCO_3 \cdot \frac{1\ mol\ NaHCO_3}{84.01 g\ NaHCO_3} \cdot \frac{1\ mol\ CO_2}{1\ mol\ NaHCO_3} \cdot \frac{58.44\ g\ NaCl}{1\ mol\ CO_2}$$

$$= 1.39\ g\ NaCl$$

3. The mass of NaCl that you ended up with is called the Actual yield because it is the amount of product that you actually got in the experiment. What was the mass of NaCl that was actually produced (see second data table)? 1.28 g NaCl

4. Calculate your % yield **(actual yield / theoretical yield) X 100**.

$$\frac{1.28}{1.39} \cdot 100 = 92.1\%$$

5. Did you get what you were supposed to? Explain one thing that might account for the difference.

No, the actual yield was slightly less than the theoretical yield, and one potential cause of this is a portion of the $NaHCO_3$ not reacting (perhaps due to it still being on the beaker).