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Virtual Lab #10: Empirical Formulas

PURPOSE:

During this virtual lab, you will learn the steps for carrying out a lab procedure to calculate the empirical formula of the product formed when magnesium ribbon burns in air. Since you know the starting mass of magnesium used and the final mass of the product formed, you can calculate the mass of oxygen from the air used in the experiment. From this data, it is possible to experimentally determine the empirical formula of magnesium oxide.

INTRODUCTION:

We can determine the atoms in a compound and the ratio of atoms combining with each other by looking at a chemical formula. For example, water has the chemical formula H_2O. This formula tells you that the compound water is formed from hydrogen and oxygen atoms combining in a 2:1 ratio.

When talking about chemical formulas, we can distinguish between empirical formulas and molecular formulas. An **empirical formula** is the simplest formula and gives you the smallest whole number ratio of atoms combining in a compound. The **molecular formula** is the true formula and tells you the actual number of each type of atom in the compound. It is possible that the empirical formula and the molecular formula of a compound may be the same, but the molecular formula is often a multiple of the empirical formula. For example, H_2O is both the empirical and the molecular formula of water, meaning that the simplest whole number ratio is 2:1 hydrogen atoms to oxygen atoms and this also represents the true formula for this compound. A compound such as glucose has an empirical formula of CH_2O (the simplest ratio) and a molecular formula of $C_6H_{12}O_6$ (the true formula.)

The molecular formula of a compound can be found if you know both the empirical formula and the molar mass of the compound. Let's use glucose as an example. The empirical formula of glucose is CH_2O, meaning that a molecule of glucose contains a carbon, hydrogen, oxygen ratio of 1:2:1. The molar mass of glucose is known to be approximately 180g/mol. To find the molecular formula of glucose when given this information, determine the molar mass of the empirical formula and compare it to the molar mass of the compound. The molar mass of CH_2O is approximately 30g/mol. If the molar mass of the molecular formula is 180g/mol, we can see that the molecular formula is 6 times larger than the empirical formula. Multiplying the subscripts in our empirical formula by 6 gives us a molecular formula of $C_6H_{12}O_6$.

Experimentally, it possible to find the empirical formula of a compound. If we know the mass of each element present in the compound, we can convert this to moles and then calculate a whole number molar ratio of one element to another. This ratio will give us the empirical formula of a compound.

Lab #4: Empirical Formulas

To experimentally calculate an empirical formula based off of data in which you know the mass of each element present in the compound, follow these steps:

1. Convert grams of each element into moles.
2. Divide moles of each element in the compound by whichever element has the smallest number of moles to obtain the mole ratio.
3. If the mole ratio is not a whole number, multiply to reach a whole number.

Here is an example:

Example 1: Solid magnesium ribbon is burned and combines with oxygen from the air to form a compound of magnesium oxide. If 0.297 grams of magnesium is ignited and 0.493 g of the oxide compound is obtained, what is the empirical formula of magnesium oxide?

The gain in mass is the mass of oxygen that combined with the magnesium:

Mass of O = 0.493 g – 0.297 g = 0.196 g

Convert grams of magnesium and grams of oxygen to moles of each compound:

Moles of Mg= $0.297\ g\ Mg \times \frac{1\ mol\ Mg}{24.3\ g\ Mg} = 0.01222\ mol\ Mg$

Moles of O = $0.196\ g\ O \times \frac{1\ mol\ O}{16.00\ g\ O} = 0.01225\ mol\ O$

We can use these mole values to get our ratio for the compound. Our formula is now: $Mg_{0.01222}O_{0.01225}$. Dividing both mole values by the smallest number of moles in the formula, we find that for this particular example the ratio is essentially a 1:1 ratio. The empirical formula of the magnesium oxide compound is thus MgO.

What if your molar ratios are not whole numbers? Say for example that you had obtained a ratio of 1.5:1. You can't have a formula $X_{1.5}Y_1$. Since subscripts must be whole numbers, it would be necessary to multiply the subscripts calculated to get whole numbers. Multiplying our subscripts by 2 would give us the formula X_3Y_2.

Lab #4: Empirical Formulas

PRELAB:

1. Write empirical formulas for the following compounds containing:

   a. 0.200 mol Al and 0.600 mole Cl

   0.007 mol Al
   0.017 mol Cl } $Al_1Cl_{2.5}$ → Al_2Cl_5

   b. 0.0400 mol Ba, 0.0400 mol S, and 0.160 mol O

   .04 mol Ba
   .04 mol S
   .160 O } $BaSO_4$

2. A student ignited lead in a crucible in an oxygen rich atmosphere to form a lead oxide compound. The mass of the empty crucible is 22.35g. The mass of the crucible and the lead before heating is 25.08g. After heading, the mass of the crucible plus lead oxide formed is 25.05g. (Show all work!)

$2Pb + 1O_2 \rightarrow 2PbO$
2:1:2
molar mass of 223.2 g
Mass of Pb.
25.08 g - 22.35g
= 2.73 g.

a.) What is the mass of Pb in the lead oxide compound?

25.08g - 22.35g = 2.73g

b.) What is the mass of O in the lead oxide compound?

28.08g - 28.05g = 0.03g

c. Calculate the moles of Pb in the lead oxide compound.

Pb mol = $2.73g \cdot \frac{1\ mol}{207.2g} = 0.0132\ mol$

d. Calculate the moles of O in the lead oxide compound.

O mol = $0.03g \cdot \frac{1\ mol}{15.99\ mol} = 0.0019\ mol$

e. Calculate the molar ratio of Pb: O and determine the empirical formula for the lead oxide compound.

Ratio of Pb:O is ≈ 7

Empirical Formula: PbO_7

3. A compound of only nitrogen and oxygen is 30.5% nitrogen. The molar mass of the molecular compound was found to be 92 g/mol. Find the empirical formula and molecular formula of the compound. (Show all work!)

30.5 % N
69.5 % O

g N = .305 · 92g/mol = 28.06 g

g O = .695 · 92g/mol = 63.94 g

N mol = $\frac{28.06\ g}{28\ g/mol} \approx 1\ mol$

O mol = $\frac{63.94\ g}{15.99\ g} = 4\ mol$

Molecular : NO_4

Empirical : NO_4

Lab #4: Empirical Formulas

EXPERIMENTAL PROCEDURE:

Go to: https://www.youtube.com/watch?v=OuFqtxZJRvM. Watch the video showing someone performing this experiment. Use the space below to write out a procedure for this experiment. Your procedure should be a numbered list of commands telling someone what steps to follow when performing this experiment:

1. Record the mass of an empty crucible (g)
2. Record mass of crucible, lid, and magnesium ribbon (g).
3. Wrap magnesium coil usely and place under lid.
4. Light bunsen burner, ensure that no flammable objects are nearby.
5. Place crucible close to bunsen burner
6. Raise lid periodically to allow O_2 in. Escaped smoke is product.
7. Leave crucible to cool once Mg no longer ignites
8. Use water and glass rod to stir into thick paste and reheat crucible to boil off water.
9. Weigh crucible once it has cooled.
10. Clean crucible before returning it

REPORT FORM: Assume that a student followed the same procedure you have just outlined above and performed the experiment. Below is given some of their experimental data. Fill in any blanks in the data using your logic, then complete.

1. Mass of empty crucible and lid. ____37.2____ g
2. Mass of empty crucible, lid and magnesium ribbon ____37.62____ g

   Mass of magnesium used 37.62 - 37.2 = ____0.42____ g
3. Mass of crucible, lid, and MgO product ____37.9____ g
4. Show all work in your empirical formula calculation!

Mass of magnesium oxide formed in experiment 37.9 - 37.2 = ____0.7____ g

Mass of oxygen formed in experiment 37.9 - 37.62 = ____0.28____ g

Moles of magnesium used: $\frac{0.7g}{24.305\ g/mol}$ = ____0.017____ mol

Moles of oxygen used: $\frac{0.28g}{15.99g/mol}$ = ____0.018____ mol

0.017 ≃ 0.018

Ratio of Mg:O = ____1:1____ Experimental Empirical formula: ____MgO____

Lab #4: Empirical Formulas

POSTLAB:

1. What are some possible sources of error in this experiment?

Some MgO product may be lost to escaping smoke and the glass rod. Additionally, some water may be left in the crucible after the burning process.

2. What is the percent magnesium in the magnesium oxide compound based on your experimental data? (Remember % composition= [mass of element/mass of compound] x 100%)

$$\% \text{ comp of Mg} = \frac{0.42g}{0.70g} \times 100 = 60\%$$

3. 0.606 g of rust (iron oxide) is formed when 0.424 g of iron is burned in oxygen rich air. Find the empirical formula of the iron oxide. What is the name for this compound? (Show all work!)

$x = m\ O$

$$4 Fe + 3 O_2 = 2 Fe_2O_3$$

$$0.424 + x = 0.606$$
$$x = 0.182$$

$$\text{mol Fe} = \frac{.424 g}{55.845 g/mol} = 0.008 \text{ mol Fe}$$

$$\text{mol O} = \frac{0.182 g}{15.99 g/mol} = 0.011 \text{ mol O}$$

$$\frac{.011}{0.008} \cong 1.5$$

Fe : O

1 : 1.5

2 : 3

Empirical formula: Fe_2O_3

Name: iron (III) oxide