

EXPERIMENT 12: Empirical Formula of a Compound

INTRODUCTION

Chemical formulas indicate the composition of compounds. A formula that gives only the simplest ratio of the relative number of atoms in a compound is the **empirical formula** or **simplest formula**. The ratio usually consists of small whole numbers. We call a formula that gives the actual numbers of each type of atom in a compound the **molecular formula**. The numbers in a molecular formula will be whole number multiples of the numbers in an empirical formula. To determine the molecular formula of a compound, we need to know both the empirical formula and the **molar mass** of the compound.

Benzene, for example, has an empirical formula of CH. In a molecule of benzene, the number of carbon atoms (C) and hydrogen atoms (H) are the same. The molar mass of benzene is 78.11 g/mol.

We can determine the molecular formula of benzene by first calculating the molar mass of the empirical formula, which is 13.02 g/mol. We then determine the number of empirical units in a molecule by dividing the molar mass of benzene by the empirical molar mass:

$$\text{Number of empirical units} = (78.11 \text{ g/mol}) / (13.02 \text{ g/mol}) = 6 \text{ empirical units}$$

Multiplying the empirical formula by 6 gives the molecular formula of benzene, (CH) × 6 or **C₆H₆**.

Experimentally, we can determine the empirical formula of a compound by first finding the **mass of each element** in a sample of the compound. We then convert the mass of each element to the equivalent number of **moles** of that element.

To find the simplest formula of a compound, we will combine the elements in the compound under conditions that allow us to determine the mass of each element. From these data, the moles of atoms of each element may be calculated. By dividing the numbers to the smallest number of moles, you obtain quotients that are in a simple ratio of integers or are readily converted to such a ratio. The ratio of moles of atoms of the elements in a compound is the same as the ratio of individual atoms that is expressed in an empirical formula.

Example: Suppose that we want to determine the empirical formula of the oxide that formed when we ignited 0.175 g of aluminum (Al) in an open container to produce a compound of Al and O that weighed 0.331 g. The gain in mass is due to the presence of oxygen atoms that combined with the aluminum atoms in the reaction.

First, we find the number of grams of each element in the sample of the compound. We know we started with 0.175 grams of Al, so we can calculate the grams of O in the compound by subtracting the mass of Al from the mass of compound:

$$\text{Mass of O} = 0.331 \text{ g} - 0.175 \text{ g} = 0.156 \text{ g}$$

Next, we find the number of moles of each element in the compound:

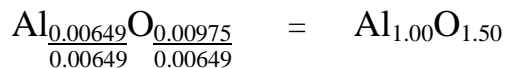
Moles of Al = (mass of Al) / (molar mass of Al) = (0.175 g) / (26.98 g/mol) = **0.00649 mol**

Moles of O = (mass of O) / (molar mass of O) = (0.156 g) / (16.00 g/mol) = **0.00975 mol**

At this stage, our formula is “Al_{0.00649}O_{0.00975}”.

Now we need to convert the subscripts to the simplest whole numbers. Use the following steps to do this:

1) Start by dividing each number of moles by smallest number of moles in the formula. The smallest number of moles is the moles of Al, 0.00649 mol.



2) Usually the subscripts at this stage are very near whole numbers, but in this case we still have a fractional subscript (1.5). To correct this, we will multiply both subscripts by a whole number (2, 3, 4, etc.) in order to obtain the simplest whole number subscripts. In this example, if we multiply 1.5 by 2, we will get a whole number, 3 for the O subscript. Therefore, multiplying both subscripts by 2, we obtain,



which gives the empirical formula Al₂O₃.

You may find it convenient to organize your calculations by arranging both data and results in a table as follows:

Element	Convert Grams to Moles	Divide Each by the Smallest Number of Moles	Convert to Whole Numbers
Al	0.175 g / 26.98 g/mol = 0.00649 mol	0.00649 mol / 0.00649 mol = 1.00	1.00 × 2 = 2.00 = 2
O	0.156 g / 16.00 g/mol = 0.00975 mol	0.00975 mol / 0.00649 mol = 1.50	1.50 × 2 = 3.00 = 3

In this experiment, you will react a known mass of magnesium (Mg) with hydrochloric acid, HCl (aq), to form a compound containing only the elements Mg and Cl (magnesium chloride). The mass of Cl reacting with the Mg will be found from the difference in the mass of the product and the mass of Mg used. By following the same sequence of calculations used in the Al₂O₃ example, you will be able to experimentally verify the empirical formula of magnesium chloride.

The reaction of magnesium with hydrochloric acid is an example of a single replacement reaction. Can you write the balanced equation for the reaction using the known formula of magnesium chloride?

Experimental Procedure

★ **Safety Note: Hydrochloric acid is a strong acid that is harmful to the skin and especially to your eyes. Wear your safety glasses or goggles during the entire procedure to protect your eyes, and avoid inhaling vapors of HCl during the drying procedure (use a fume hood if possible). The reaction also produces flammable hydrogen gas (H₂), so Bunsen burners should not be used while the reaction is in progress.**


1) Obtain a clean, dry evaporating dish and weigh it to the nearest 0.01 g. Record this value on the report form.



2) Place a small piece of magnesium ribbon into the evaporating dish and record the mass on the report form. From the difference in masses, record the mass of magnesium used. Note: The mass of magnesium should not exceed 0.15 grams, or the product may be difficult to dry (see below).

3) Measure 3 mL of 6 M HCl* in a 10 mL graduated cylinder (or you may use a pipet) and carefully add the HCl solution to the evaporating dish containing the Mg ribbon (Caution: Vigorous reaction!). Allow the reaction to proceed until the reaction is complete, giving a clear solution with no magnesium particles remaining.

4) Place the evaporating dish on an electric hot plate and heat to nearly boiling. Avoid excessive heat that can cause dangerous splattering of hot HCl!

5) Heat the solution until evaporation of the water is complete. The white, solid product that remains is magnesium chloride. It is difficult to tell by appearance when the product is completely dry, so we will use a method called heating to constant weight. When the product appears thoroughly dry, carefully remove the hot evaporating dish from the hot plate (you may use crucible tongs  to handle the hot dish), allow it to cool, and record the weight on the report sheet under “first weighing.” Then, place the dish back on the hot plate and heat for 10 additional minutes. Allow the dish to cool and record the weight under “second weighing.” If the second weighing agrees with the first weighing, you may reasonably assume that drying is complete. If the second weighing is less than the first weighing, place the dish back on the hot plate and heat for 10 more minutes and obtain a third weighing. Repeat this process until successive weighings agree to within 0.01 g.

An evaporating dish is much better container than a beaker for this process for two reasons:

- The solid product spreads out more in the evaporating dish as the liquid evaporates, facilitating drying.
- In a beaker, the vapors of liquid tend to condense on the walls and run back down, inhibiting drying.

6) From your data, calculate the moles of Mg and the moles of Cl in the product. From the number of moles of each element, determine the empirical formula of magnesium chloride.

7) Wash your evaporating dish with water (the product may be washed down the sink) and return your equipment to their proper storage locations before leaving the lab.

* The “M” in “6 M HCl” stands for molarity. This is a standard concentration unit in chemistry, and it means moles of solute per liter of solution, mol/L. One liter of 6 M HCl solution contains 6 moles of HCl.

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REPORT FORM

Name _____

Instructor _____

Date _____

1. Mass of empty evaporating dish _____ g
2. Mass of evaporating dish and magnesium _____ g
3. Mass of magnesium [2] – [1] _____ g
4. Mass of evaporating dish and magnesium chloride (after heating and cooling)
 - First weighing _____ g
 - Second weighing _____ g
 - (if necessary) Third weighing _____ g
 - (if necessary) Fourth weighing _____ g
5. Mass of magnesium chloride [4] – [1] _____ g
6. Mass of chlorine in magnesium chloride [5] – [3] _____ g
7. Moles of magnesium (show your calculation) _____ mol
8. Moles of chlorine (show your calculation) _____ mol
9. Moles of magnesium divided by the smaller number of moles (3 sig. figures) (show your calculation) _____
10. Moles of chlorine divided by the smaller number of moles (3 sig. figures) (show your calculation) _____
11. Your experimental empirical formula of magnesium chloride (with whole number subscripts) _____
12. True (known) empirical formula of magnesium chloride _____

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Pre-Laboratory Questions and Exercises

Due before lab begins. Answer in the space provided.

1. a) How many moles of copper atoms are in 150 g of copper metal?

b) How many copper atoms are in this amount of copper?

2. Write the empirical formula for the following compounds containing,
 - a) 0.0200 mole of Al and 0.0600 mole of Cl.

 - b) 0.0800 mole of Ba, 0.0800 mole of S, and 0.320 mole of O.

3. When 0.424 g of iron powder is burned in an oxygen atmosphere, 0.606 g of a reddish brown oxide is obtained. Determine the empirical formula of the oxide.

4. What safety precautions are cited in this experiment?

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Post-Laboratory Questions and Exercises

Due after completing the lab. Answer in the space provided.

1. Why was an evaporating dish more suitable for this lab procedure, rather than using a beaker?
2. How would your experimental formula of magnesium chloride " MgCl_x " have been affected if your product was not dried completely before weighing it? Would "x" be too high or too low?
3. When 6.25 grams of pure iron are allowed to react with oxygen, a black oxide forms. If the product weighs 8.15 g, what is the empirical formula of the oxide?
4. A compound of nitrogen and oxygen is 30.46% by mass N and 69.54% by mass O. The molar mass of the compound was determined to be 92 g/mol.
 - a) What is the empirical formula of the compound?
 - b) What is the molecular formula of the compound?