

Formulas and Percent Compositions of Ionic Compounds

Strand	Molar Relationships
Topic	Investigating stoichiometry
Primary SOL	CH.4 The student will investigate and understand that chemical quantities are based on molar relationships. Key concepts include b) stoichiometric relationships.
Related SOL	CH.3 The student will investigate and understand how conservation of energy and matter is expressed in chemical formulas and balanced equations. Key concepts include c) writing chemical formulas.

Background Information

Percent composition: To determine the percent composition (percent by mass) of an ionic compound, calculate the molar mass of the compound by writing its formula and adding up the masses of each element in it. Calculate the percent of each element in the compound (the percent composition) by dividing the mass of each element by the molar mass of the entire compound and then multiplying by 100. As an example, for ammonium sulfate, $(\text{NH}_4)_2\text{SO}_4$, the steps are as follows:

- Molar mass is 132.154 g/mol
- Percent H = $8(1.008 \text{ g}) \div 132.145 \text{ g} = 0.061 \times 100 = 6.10\% \text{ H}$
- Percent N = $2(14 \text{ g}) \div 132.145 = 0.212 \times 100 = 21.2\% \text{ N}$
- Percent S = $1(32 \text{ g}) \div 132.145 = 0.242 \times 100 = 24.2\% \text{ S}$
- Percent O = $4(16 \text{ g}) \div 132.145 = 0.484 \times 100 = 48.5\% \text{ O}$
[or percent O = $100 - (6.10 + 21.2 + 24.2) = 48.5\% \text{ O}$]

Empirical formula determination: One of the most important things we can learn about a compound is its chemical formula. The formula for a compound that expresses the smallest whole-number ratio of the atoms present is called an “empirical formula.” For example, a compound with the formula $\text{C}_4\text{H}_8\text{O}_4$ has the same empirical formula as a compound with formula $\text{C}_6\text{H}_{12}\text{O}_6$: the empirical formula for both of these compounds is CH_2O .

The formula for a molecular compound that gives the actual number of atoms per molecule is called a “molecular formula.” The molecular formula $\text{C}_6\text{H}_{12}\text{O}_6$ can be represented as a whole-number multiple of the empirical formula: $(\text{CH}_2\text{O})_6$.

To calculate the empirical formula for a compound, we first determine the relative masses of the various elements that are present. One way to do this is to measure the masses of elements that react to form the compound. For example, suppose we weigh out 0.2636 g of pure nickel metal into a crucible and heat the metal in air so that the nickel can react with oxygen to form a nickel oxide compound. After the sample has cooled, we weigh it again and find its mass to be 0.3354 g. The gain in mass is due to the oxygen that reacts with the nickel to form the oxide. Therefore, the mass of oxygen present in the compound is the total mass of the product minus the mass of the

nickel: $0.3354 \text{ g} - 0.2636 \text{ g} = 0.0718 \text{ g}$. Note that the mass of nickel present in the compound remains the same as the nickel metal originally weighed out. So we know that the nickel compound contains 0.2636 g nickel and 0.0718 g oxygen.

To determine the empirical formula for a compound,

- convert the grams of each element in the compound to moles of atoms of each element
- determine the whole-number mole ratio of each element to the other(s)
- write the formula.

For the above example involving nickel, the steps are as follows:

- $0.2636 \text{ g Ni} \div 58.71 \text{ g per mol Ni atoms} = 0.00449 \text{ mol Ni atoms}$
- $0.0718 \text{ g O} \div 16.00 \text{ g per mol O atoms} = 0.00449 \text{ mol O atoms}$

These mole quantities contain an equal number of atoms. It is clear from the moles of atoms of each substance that the mole ratio is 1:1. Therefore, the formula is NiO. This is the empirical formula, expressing the smallest whole-number ratio of atoms. All ionic compounds are expressed as empirical formulas—that is, we would never write, for example, Ni_2O_2 .

For a compound containing 1.3813 g of Pb, 0.00672 g H, 0.4995 g As, and 0.4267 g O, the steps to find the empirical formula for the compound are as follows:

- $1.3813 \text{ g Pb} \div 207.2 \text{ g per mol Pb} = 0.006667 \text{ mol Pb}$
- $0.00672 \text{ g H} \div 1.008 \text{ g per mol H} = 0.006667 \text{ mol H}$
- $0.4995 \text{ g As} \div 74.92 \text{ g per mol As} = 0.006667 \text{ mol As}$
- $0.4267 \text{ g O} \div 16.00 \text{ g per mol O} = 0.026669 \text{ mol O}$

Divide these mole values by the smallest number of moles to get the whole-number mole ratio:

- $0.006667 \text{ mol Pb} \div 0.006667 = 1 \text{ mol Pb}$
- $0.006667 \text{ mol H} \div 0.006667 = 1 \text{ mol H}$
- $0.006667 \text{ mol As} \div 0.006667 = 1 \text{ mol As}$
- $0.026669 \text{ mol O} \div 0.006667 = 4 \text{ mol O}$

Therefore, the empirical formula for the compound is PbHAsO_4 .

When a 0.3546 g sample of vanadium metal is heated in air, it reacts with oxygen to form an oxide compound with a mass of 0.6330 g . To calculate the empirical formula for this sample of vanadium oxide, the following steps are used:

- $0.6330 \text{ g} - 0.3546 \text{ g} = 0.2784 \text{ g}$ (mass of oxygen that reacted)
- $0.3546 \text{ g V} \div 50.94 \text{ g per mol V} = 0.006961 \text{ mol V}$
- $0.2784 \text{ g O} \div 16.00 \text{ g per mol O} = 0.01740 \text{ mol O}$

Divide these mole values by the smallest number of moles to get the whole-number mole ratio:

- $0.006961 \text{ mol V} \div 0.006961 = 1.000 \text{ mol V}$
- $0.01740 \text{ mol O} \div 0.006961 = 2.500 \text{ mol O}$

Since one of these numbers (2.500) is not a whole number, multiply both numbers by the lowest possible number (2) that will result in two whole numbers and therefore form a whole-number ratio:

- $1.000 \times 2 = 2 \text{ V}$
- $2.500 \times 2 = 5 \text{ O}$

Hence, the whole-number ratio is 2:5. (Note that this step is necessary only if the mole numbers are not already whole numbers.) Therefore, the empirical formula for this sample of vanadium oxide is V_2O_5 .

Materials

- Standard test tubes (18 × 150 mm)
- Wooden splints
- M HCl
- Beral pipettes
- Magnesium ribbon
- Safety goggles

Student/Teacher Actions (what students and teachers should be doing to facilitate learning)

Introduction

1. Introduce the activity by telling students that they will do a lab in which they will form an ionic compound from the elements magnesium and chlorine. They will determine the empirical formula for this compound by comparing the moles of each element that react and determining the mole ratio of one element to the other. They will also determine the percent composition of the compound.

Procedure

Have students do the experiment, as follows:

1. Make a data chart with spaces for the following data:
 - Mass of clean, dry test tube
 - Mass of magnesium ribbon
 - Mass of test tube and product
 - Mass of product
 - Observations of gas splint test
 - Identity of gas
2. Select a standard 18 × 150 mm test tube, and make sure it is clean and dry. Find the mass of the test tube, and record it in the data chart.
3. Find the mass of approximately 0.10 grams of magnesium ribbon. It is not necessary to have exactly 0.10 grams, but do not be too far off. Record in the data chart the exact mass of the magnesium ribbon used.
4. Fill a beral pipette with 3.0 M HCl. **CAUTION! Strong acid.** Add the acid to the test tube, a few drops at a time. Run the drops down the inside wall of the test tube, holding the test tube at a gentle angle. **CAUTION! Do not add the acid all at once.**
5. While the reaction is proceeding, test the gas with a burning splint. What is the identity of the gas?
6. Continue to add the acid until the reaction is complete—i.e., when there is no solid magnesium visible. **CAUTION! Try to avoid adding excess acid.**
7. Set up a Bunsen burner. Using a *small* flame, start to evaporate off the liquid in the test tube. Be sure to hold the test tube at a 45-degree angle. Keep the test tube moving in the

flame, and do not heat just the bottom of the test tube. **CAUTION! Do not point the test tube at anyone, including yourself.**

8. When evaporation appears complete, test by removing the test tube from the flame and inverting a *clean, dry test tube* over the mouth of your test tube. If you observe condensation forming in the top test tube, then continue heating the original test tube a few minutes longer and test again with another clean, dry test tube. Be patient; the heating process will probably take about 10 to 15 minutes.
9. When you are sure your product is dry, let your test tube cool completely. Find the mass of the test tube and product, and record this mass in the data chart.
10. When you are finished, clean up. You may wash your product down the sink.

Observations and Conclusions

1. Prompt conclusions by having students do the following:
 - Calculate the number of moles of magnesium in the compound.
 - Calculate the mass of chlorine that reacted with the magnesium.
 - Calculate the number of moles of chlorine in the compound.
 - Find the whole-number mole ratio of Mg to Cl in the compound. (Dividing the number of moles of each element by the smaller number of moles gives the whole-number mole ratio.)
 - Write the formula for the compound, and compare it to the known formula MgCl_2 .
 - Calculate the percent Mg and the percent Cl in the compound from both your lab data and the known chemical formula MgCl_2 .
 - Explain how the percent compositions compare with each other
 - Explain some possible causes of error in your experiment.

Assessment

- **Other**
 - Use students' lab observations and calculations for assessment.
 - Use the questions on the attached handout for assessment. Have students work individually or with their lab partners.

Extensions and Connections (for all students)

- Have students find the empirical formulas for additional ionic compounds, such as ZnCl_2 and AlCl_3 , using the same basic procedures.

Finding the Formula and Percent Composition of an Ionic Compound

Directions: Answer the following questions. Show all your calculations, rounding your answers to the teacher-specified number of significant digits and labeling units clearly.

1. What is the percent by mass of each element in the compound $C_{14}H_{18}N_2O_5$?
2. What is the percent composition of sucrose ($C_{12}H_{22}O_{11}$)?
3. What is the percent of phosphorous in $Sn_3(PO_4)_2$?
4. What is the percent composition of aluminum oxide?
5. What is the percent composition of heptaphosphorous decoxide?
6. What is the percent by mass of iron in iron(III) chloride?
7. An oxide of aluminum metal is formed by the reaction of 4.151 grams of Al with 3.692 grams of O. What is the empirical formula for this compound?
8. A chemist analyzed an unknown compound and collected the following data: 0.8007 g C, 0.9333 g N, 0.2016 g H, 2.133 g O. What is the empirical formula for the compound? What is the percent composition of the compound?
9. In a lab experiment, it was observed that 0.6884 grams of lead combines with 0.2356 grams of oxygen. What is the empirical formula for this compound?
10. Cisplatin is used to treat cancerous tumors. It has the following composition: 65.02 g Pt, 9.34 g N, 2.02 g H, and 23.63 g Cl. What is the empirical formula for cisplatin?
11. The most common form of nylon is 63.68 g C, 12.38 g N, 9.80 g H, and 14.14 g O. What is the empirical formula for this common form of nylon?
12. If 1.000 g of barium metal is heated in a stream of pure oxygen gas, 1.117 grams of an oxide is produced. What is the empirical formula for this oxide?
13. A compound has the following composition: 58.84 g Ba, 13.74 g S, and 27.43 g O. What is the empirical formula for this compound?
14. If cobalt metal is mixed with sulfur and heated strongly, 100 grams of a compound is produced that contains 55.06 grams of cobalt. What is the empirical formula for the compound?
15. When 3.269 grams of zinc metal is heated in pure oxygen, the sample gains 0.800 grams of oxygen in forming the oxide. What is the empirical formula for the oxide?
16. A compound consists of 65.45 g C and 5.492 g H and 29.06 g O. What is the empirical formula for the compound?
17. Phosphorous forms two oxides. One has 56.34 g P and 43.66 g O. The other has 43.64 g P and 56.36 g O. What are the empirical formulas for these two oxides?
18. Magnetite is an iron ore. 100 g contains 72.4 g Fe and 27.6 g O. What is its empirical formula?
19. What is the percentage of water by mass in the compound $CoCl_3 \cdot 6H_2O$?
20. A compound with 26.2 g N, 7.5 g H, and 66.3 g Cl is sometimes used to salt icy roads. What is the empirical formula for this compound?