

Experiment 3
The Empirical Formula of a Copper Oxide

Reading assignment: Chang, Chemistry 9th edition, pp. 86-92.

Goals

The reaction of hydrogen gas with a copper oxide compound will be studied quantitatively. By measuring the masses of copper and oxygen in copper oxide we will determine its empirical formula.

Equipment and Materials

Solid copper oxide, zinc metal in a hydrogen-generator bottle with thistle tube, 3 M sulfuric acid, a large test tube, a small test tube, two beakers (100 mL), Bunsen burner, calcium sulfate in a drying tube.

Safety Note: Safety glasses are required when performing this experiment

Discussion

Copper was perhaps the first metal to be produced and used by humans. A pendant made of copper discovered in Northern Iraq may have been produced over 10,000 years ago. Native cultures in the American mid-west used copper 6,000 years ago. Copper mines are known to have existed in Egypt at least 3,000 years ago.

Copper is an extremely important material in modern society. The metal has high electrical and thermal conductivities, is resistant to corrosion (oxidation), and can be mixed with other metals to form alloys. Brass is formed from a mixture of copper and zinc. Bronze is formed from copper and tin. In the United States copper is used at a rate of 7.3 billion pounds (3.3 million metric tons) per year.¹

Table 1: Major uses of copper.¹

Building construction	46%
Electrical/electronic products	23%
Consumer products	11%
Transportation products	10%
Industrial machinery	10%

Owing to its importance in the modern world, copper is mined in large amounts throughout the world. Large copper mining industries exist around the world.

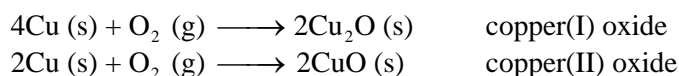
Table 2: Annual mining production of copper in megatons.²

Nation	Amount	Nation	Amount
Chile	2.3	Indonesia	0.38
United States	1.89	Peru	0.35
Canada	0.74	Zambia	0.35
Russia	0.60	Poland	0.34
Australia	0.42	Kazakhstan	0.22
Peru	0.40	Philippines	0.11

The environmental effects of copper mining are significant. Toxic metals can be released into fresh water sources in the process of mining.³ Copper is a valuable enough resource that it is recycled to a large degree. Approximately 40% of copper is recycled in the United States.

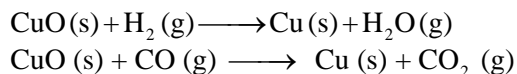
Oxidation States of Copper

Two forms of copper oxide are found in nature, copper(I) oxide and copper(II) oxide. Each compound can be formed from the reaction of copper metal with molecular oxygen:

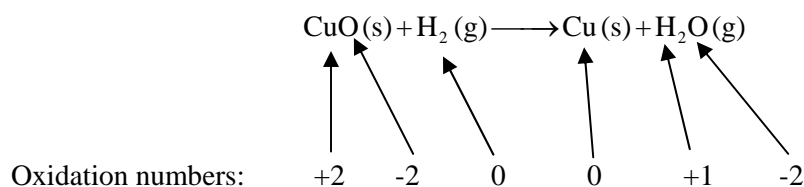


Copper(I) oxide, which is sometimes called cuprous oxide, is a deep red (almost black) material that is used commercially in fungicides and insecticides. It is found in nature as a mineral called cuprite.

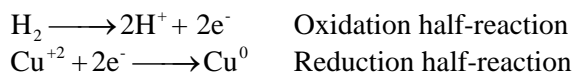
Copper(II) oxide is sometimes called cupric oxide and is a dark black material. Copper(II) oxide can be reduced to metallic copper by reaction with molecular hydrogen or gaseous carbon monoxide at high temperature:



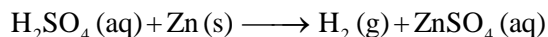
In both reactions the products include metallic copper. In this experiment we will use the first reaction (copper (II) oxide with hydrogen) in order to determine the empirical formula of copper oxide. In this reaction hydrogen is oxidized to H^+ and copper (II) is reduced to Cu^0 . This can be shown by considering the oxidation numbers for copper and hydrogen in the equation:



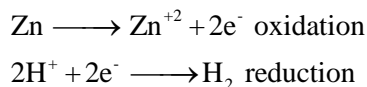
In the equation above the oxidation number of copper has changed from +2 to 0. The oxidation number of hydrogen has changed from 0 to +1. In half-reaction notation:



The reduction of copper(II) to copper requires that a reducing agent (hydrogen or carbon monoxide) be added. Hydrogen gas can be produced by adding a strong acid like sulfuric acid to zinc metal. The reaction is fast if the concentration of sulfuric acid is high and the surface area of the metallic zinc is large. The reaction is another example of reduction-oxidation:



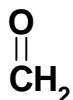
The zinc metal (Zn) is oxidized to Zn^{+2} and H^+ is reduced to H_2 :



The hydrogen gas formed in this way can then be used to reduce copper oxide to metallic copper.

Empirical Formulas

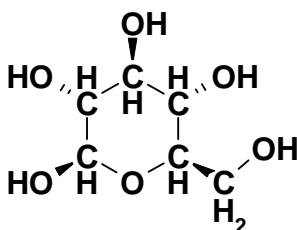
The empirical formula of a compound is the simplest whole number ratio of atoms in the compound. For example, the empirical formula for glucose is CH_2O . This means that there is 1 carbon atom per 2 hydrogen atoms per 1 oxygen atom in a molecule of glucose. The empirical formula is useful because it can be used to determine the molecular formula of a compound. The molecular formula gives the actual number of atoms in one molecule of a compound. For glucose, the molecular formula is $\text{C}_6\text{H}_{12}\text{O}_6$. If the subscripts of the molecular formula for glucose are divided by 6 then the result is the empirical formula (CH_2O).



formaldehyde

Empirical Formula: CH_2O

Molecular Formula: CH_2O



glucose

Empirical Formula: CH_2O

Molecular Formula: $\text{C}_6\text{H}_{12}\text{O}_6$

To determine the empirical formula of a compound we need to know the mass of each element in a sample of the compound.

Sample Calculation

Suppose a sample of copper oxide weighs 1.5000 g and is reacted with hydrogen at high temperature to form copper and water. After heating the sample, the mass of the product is measured at 1.1955 g. The mass of the product (assuming the water evaporates) is solely due to the copper product. We also know the mass of the copper(II) oxide. The mass of the copper oxide is equal to the masses of the copper and the oxygen that comprise the material. So the mass of the oxygen is equal to the difference between the mass of the copper(II) oxide and the mass of the copper that is produced. Since we know the mass of the copper and the oxygen we can find the number of moles of copper and oxygen from their atomic masses. Once the moles of each element are known we can determine the empirical formula by dividing the number of mole for each element by the smallest number of moles.

mass of oxygen = mass of copper oxide - mass of copper = 1.5000 g - 1.1955 g = 0.3045 g oxygen

$$\text{moles of copper} = (1.1955 \text{ g Cu}) \left(\frac{1 \text{ mol Cu}}{63.546 \text{ g Cu}} \right) = 0.018881 \text{ mol Cu}$$

$$\text{moles of oxygen} = (0.3045 \text{ g O}) \left(\frac{1 \text{ mol O}}{16.00 \text{ g O}} \right) = 0.01903 \text{ mol O}$$

$$\text{Divide by smallest number: Copper} = \frac{0.01881 \text{ mol Cu}}{0.01881 \text{ mol Cu}} = 1.000 \quad \text{Oxygen} = \frac{0.01903 \text{ mol O}}{0.01881 \text{ mol O}} = 1.012 \approx 1$$

Each ratio is either equal to or close to 1. So, Cu_1O_1 is the empirical formula. By convention, 1 is not shown in chemical formulas. So the empirical formula for the copper oxide compound is CuO (copper(II) oxide).

References

1. Copper Development Association of America. For the year 2002.
2. Report from the Mineral Commodity Summary. 1995.
3. U.S. Environmental Protection Agency Report. May 2000.

Procedure

SAFETY PRECAUTIONS

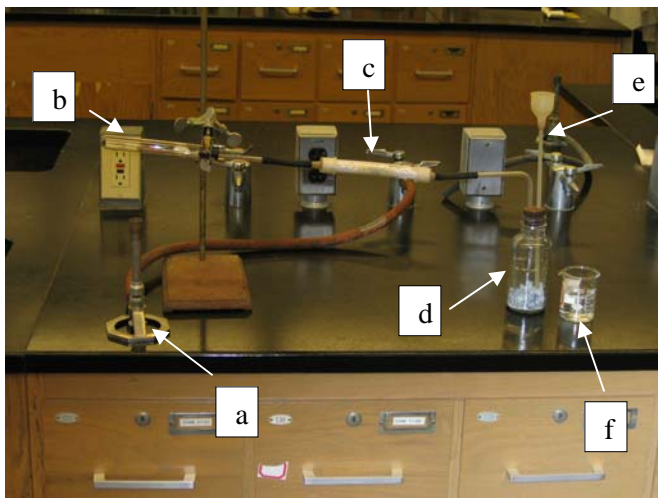
Safety glasses or goggles are required for this experiment. Sulfuric acid can burn skin and should be handled with care. Hydrogen gas is explosive and must be handled carefully. Flames from the Bunsen burner must be kept away from combustible material like paper and clothing.

Students work in pairs.

1. Use a digital analytical balance to measure the mass of a clean dry test tube provided by your instructor.
2. Add 1.300 g - 2.000 g of copper oxide powder to the test tube and re-measure the mass of the tube.
3. Obtain about 60 mL of sulfuric acid solution in a 100 mL beaker.
4. Set up the hydrogen generator apparatus described below.

The figure to the right shows the experimental apparatus used to reduce copper oxide. The important components are:

- a. Bunsen burner with gas hose.
- b. Test tube containing copper oxide.
- c. Drying tube.
- d. Hydrogen generator containing zinc.
- e. Thistle tube and funnel for transferring sulfuric acid.
- f. 100 mL beaker containing sulfuric acid solution.



The hydrogen gas generator with thistle tube is used to produce hydrogen gas. Sulfuric acid is poured into the funnel and down the thistle tube. Use a pipe cleaner, if needed, to remove water in the system between the drying tube and the nozzle tip. All connections must be firm. The test tube must be clamped right at the mouth and must have a very slight tilt downward at the mouth to allow water to flow away from the copper oxide. The copper oxide must be confined to the rear half of the tube and not be touched by the long tubing.

Generating Hydrogen Gas

1. Ignite the Bunsen burner. Do not start heating the sample yet.
2. Slowly pour about one-half to two-thirds of the 60 mL of 3 M sulfuric acid (H_2SO_4) into the thistle tube. As the sulfuric acid oxidizes the zinc metal there should be a bubbling gas visible in the generator bottle. The sulfuric acid should be added slowly so that the mixture does not bubble out through the

thistle tube. The reaction produces a large amount of heat and can boil the water in the sulfuric acid solution, so adding slowly will decrease the likelihood that the solution will bubble out of the funnel.

3. The glass tubing that is connected to the test tube is fitted with an opening that allows excess hydrogen to flow out of the system. This hydrogen flowing out of the system can be ignited and burned by collecting some of it with a small test tube and then igniting the collected gas with the Bunsen burner. If the test tube is quickly returned to the overflow opening then the heat generated can also ignite the hydrogen flowing out of the opening.

4. Continue adding the sulfuric acid into the thistle tube as the reaction proceeds.

5. Place an upside down test tube on the nozzle for about 5-10 seconds. Then quickly remove the tube and bring it to the flame and place the tube's opening to the edge of the flame. There may be a soft or loud popping sound. If so, return the tube to the overflow tube and ignite the hydrogen flowing out of the opening.

6. Do **not** ignite the hydrogen with a burner or match. If after several tries you have not succeeded in igniting the hydrogen, ask for help from your instructor.

7. Place a dry upside down beaker (100 mL or 250 mL from your drawer) over the burning hydrogen flame. You should see the condensation of water on the inside surface of the beaker.

Reaction of Hydrogen with Copper(II) Oxide

1. Heat the tube containing the copper oxide gently with the Bunsen burner.

2. After a few minutes of gentle heating, apply a strong and hot flame to one part of the oxide for 5 minutes, then move the flame to another part of the oxide for an additional 10 minutes, repeating until all the oxide has had a chance to react completely with hydrogen.

3. Heating the oxide may cause the hydrogen flame to extinguish. There is no need to re-ignite it.

4. After a total of 15-20 minutes of heating the oxide should be red, indicating that the material has been reduced to metallic copper.

5. Turn off the Bunsen burner and allow the test tube to become cool to the touch while continuing the flow of hydrogen.

6. Keep the test tube clamped to the ring stand while you detach the stopper and tubing unit so that no liquid runs back into the solid powder.

7. Use a wadded up paper towel to absorb all the water in the tube.

8. Measure the mass of the test tube on the analytical balance.

9. Do not discard the test tube contents until you have shown the data to the instructor.

10. Disconnect the generator bottle and fill it with water to stop the hydrogen evolution. Pour the liquid into the sink. Do not allow the zinc metal to fall into the sink since it can be re-used.

Calculating the Empirical Formula of Copper(II) Oxide

1. Copper oxide is composed of copper and oxygen. Calculate the masses of copper oxide, copper, and oxygen from the mass measurements on the data sheet. The mass of the copper is measured after the reaction is complete. The mass of oxygen is the difference between the mass of CuO and the mass of the copper.

2. Use the atomic mass of copper to determine the number of moles of copper present in the copper oxide sample:

$$\text{moles Cu} = (\text{mass of Cu}) \left(\frac{1 \text{ mol Cu}}{63.546 \text{ g Cu}} \right)$$

3. Use the atomic mass of oxygen to determine the number of moles of oxygen present in the copper oxide sample.

$$\text{moles O} = (\text{mass of O}) \left(\frac{1 \text{ mol O}}{16.00 \text{ g O}} \right)$$

4. The empirical formula is the smallest whole number ratio of copper to oxygen atoms. To obtain the empirical formula divide the number of moles of both copper and oxygen by the smaller of the two numbers.

5. After dividing by the smaller of the two numbers round each to the nearest whole number.

Questions

1. Write balanced equations for the three main reactions that took place in this experiment:

- The preparation of hydrogen from zinc and sulfuric acid.
- The reaction of hydrogen gas with oxygen gas.
- The reaction of hydrogen with copper(II) oxide.

2. What additional information is needed to determine the molecular formula from the empirical formula?

3. Copper has two natural isotopes: copper-63 and copper-65. The atomic masses and natural abundances for each isotope are shown below. What is the weighted mass average for copper?

Copper-63	62.929599u	69.17%
Copper-65	64.927792u	30.83%

4. How many protons, neutrons, and electrons are there in each isotope of copper? Write the isotopic symbols for each isotope.

5. There are three naturally-occurring isotopes of oxygen, oxygen-16, oxygen-17, and oxygen-18. Given the data below for these three isotopes of oxygen, calculate the weighted mass average for oxygen.

Isotope	Natural abundance (%)	Mass (u)
oxygen-16	99.76	15.99491
oxygen-17	0.04	16.99913
oxygen-18	0.2	17.99916

6. How many protons, neutrons, and electrons are there in each isotope of oxygen? Write the isotopic symbols for each isotope.
7. A sample is found to be composed of 1.250 g Cu, 1.259 g O, and 0.631 g S. Calculate the empirical formula for this compound.
8. A sample is composed of 3.736 g K, 5.250 g Mn, and 6.115 g O. What is the empirical formula of the compound?
9. A compound has the following composition: 40.27% potassium, 26.78% chromium, and 32.96% oxygen. What is the empirical formula of this compound?
10. A compound has the following composition: 30.88% sodium, 21.49% oxygen, and 47.62% chlorine. What is the empirical formula of this compound?
11. Submit a one-paged (typed) report that discusses the following topics:
 - a. What were the goals or objectives of the experiment
 - b. What were your results as they relate to the goals and objectives.
 - c. Discuss the quality of your results.

Observations and Notes
Experiment 3: The Empirical Formula of a Copper Oxide

Date _____

Data and Report Sheet
Experiment 3: Empirical Formula of a Copper Oxide

Name _____ Section _____ Date _____

Partner's Name _____

1. Mass of empty test tube _____ g
2. Mass of copper oxide + test tube _____ g
3. Mass of copper + test tube _____ g
4. Mass of copper oxide _____ g
5. Mass of copper _____ g
6. Mass of oxygen _____ g
7. Moles of copper (use atomic mass of Cu) _____ mol
8. Moles of oxygen (use atomic mass of O) _____ mol
9. Empirical formula of copper oxide _____

