

Chemistry 1210 Experiment 4: The Law of Conservation of Mass

Objective

In this experiment you will burn a sample of magnesium metal in air and demonstrate the increase in mass due to formation of magnesium oxide, MgO. Then using the reaction stoichiometry you will verify the Law of Conservation of Mass.

Introduction

Antoine Lavoisier (1743-1794) is heralded as one of the fathers of modern day chemistry partially because he was the first scientist to propose the Law of Conservation of Mass. Simply put; this law states that the total mass of products resulting from a chemical reaction can be neither more nor less than the total mass of the starting reactants. In other words, mass is neither gained nor lost during a chemical reaction. This concept may seem intuitive to us now, but was certainly not obvious in the days of alchemists who believed that matter could be transmuted from one form into another (i.e. lead transmuted into gold for obvious reasons). Consider the following two examples. When mercury (II) oxide, HgO, is heated it decomposes to form mercury metal and oxygen gas. If the emission of the colorless and odorless oxygen gas is not detected, the HgO appears to lose mass on heating. Alternatively, many metals, including magnesium, increase in mass when heated because they combine with oxygen from the air to form a metal oxide. It was Antoine Lavoisier, using remarkable accurate balances, who carried out a series of carefully designed experiments in closed systems, who conclusively demonstrated the Law of Conservation of Mass.

In today's experiment magnesium metal, Mg, will be heated in a porcelain crucible where it will undergo a combination reaction with oxygen gas, O₂, from the air to form Magnesium oxide, MgO as described by Equation 1.



At the end of the experiment the crucible and its contents will have a greater mass because of the addition of one oxygen atom for each magnesium atom in the original sample. This gain in mass will be measured as accurately as possible.

One formula unit of MgO contains one Mg⁺² cation and one O⁻² anion. It is important to note that this 1:1 ratio is not a mass ratio. Since magnesium cations have a different, heavier, mass than oxygen anions, the cation:anion ratio is different than the mass:mass ratio. To further illustrate this principle consider that every automobile has a steering wheel so there is a 1:1 ratio between automobiles and steering wheels. However, the steering wheel has a much smaller mass than the rest of the automobile so there is not a 1:1 correspondence between their masses.

In the case of atoms and ions involved in a chemical reaction we determine the molar ratio between the reactants involved in a reaction based on the coefficients in the balanced chemical equation. In our example in Equation 1, two moles of magnesium react with one mole of oxygen to form two moles of MgO. In this example two moles of magnesium are required because in nature oxygen is a diatomic molecule not a single oxygen atom. This use of a balanced chemical reaction to determine the ratios of the masses of reactants to products in a chemical reaction is called stoichiometry. The mass

ratio of Mg to O₂ can be calculated based on Equation 1 and the molar masses of the two reactants using dimensional analysis and stoichiometry as described in Equations 2-5.

Let's assume that you wanted to calculate the mass of O₂ from the atmosphere that would be required to completely react with 2.50 g of Mg. You would need to go through the following process:

$$(2.50 \text{ g Mg})(1 \text{ mole Mg}/24.31\text{g}) = 0.103 \text{ moles of Mg} \quad \text{Equation 2}$$

Then, since two moles of Mg react with one mole of O₂

$$(0.103 \text{ moles Mg})(1 \text{ mole O}_2/2 \text{ mol Mg}) = 0.0515 \text{ mol O}_2 \quad \text{Equation 3}$$

The mass of O₂ that will react with 2.50 g of Mg will be

$$(0.0515 \text{ mol O}_2)(32.0\text{g O}_2/\text{mole}) = 1.65 \text{ g O}_2 \quad \text{Equation 4}$$

The mass of MgO product that will form is

$$1.65 \text{ g O}_2 + 2.50 \text{ g Mg} = 4.15 \text{ g MgO} \quad \text{Equation 5}$$

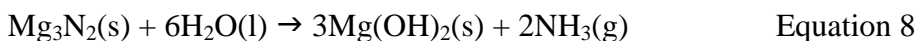
Note that stoichiometry also allows us to directly calculate the mass of MgO formed.

$$(2.50 \text{ g Mg})(1 \text{ mol Mg}/24.31 \text{ g})(2 \text{ mol MgO}/2 \text{ mol Mg})(40.31\text{g MgO}/\text{mol}) = 4.15 \text{ g MgO} \quad \text{Equation 6}$$

Equation 1 is a deceptively simple summary of the reaction of magnesium burning in air. Magnesium burns at such a high temperature that it also reacts with nitrogen gas, N₂, in the air to form black magnesium nitride, Mg₃N₂ (Equation 7).



The product of burning magnesium in air is, at this stage, a mixture of magnesium oxide and magnesium nitride. However, Mg₃N₂ readily decomposes upon addition of water to form insoluble magnesium hydroxide Mg(OH)₂ and gaseous ammonia, NH₃ (Equation 8).



The magnesium hydroxide produced in equation 8 decomposes on heating to magnesium oxide and water. Therefore, at the end of a carefully executed experiment the only product remaining in the crucible will be MgO. However, since MgO and Mg(OH)₂ are both white solids, a visual inspection cannot tell whether all the Mg(OH)₂ has decomposed. It is therefore necessary to reheat the sample and reweigh it, repeating until a constant mass is obtained. This heating and reweighing process may need to be repeated several times to obtain a constant mass.

Procedure

1. Attach a medium ring to a ring stand. Place a ceramic triangle on the ring and a Bunsen burner under the ring.
2. Obtain a clean, dry crucible and lid. Place the crucible with its lid in place in the triangle and heat it with a hot flame for two minutes. Using crucible tongs remove the crucible from the triangle and place it on wire gauze to cool. Be careful not to drop or contaminate the crucible. Always transfer the crucible and its lid separately. From this point on do not handle the crucible or the lid with your fingers.
3. When the crucible is cool, weigh it and record the mass on your data sheet.
4. Obtain a strip of magnesium metal. Sand it with a piece of sand paper or steel wool to remove the surface oxide layer. Either cut the strip into small pieces directly into the crucible or use tweezers to roll it into a coil and place the coil in the crucible.
5. Reweigh the crucible with the magnesium. Record the mass on your data sheet.
6. Place the crucible and its contents (without the lid) into the ceramic triangle over the Bunsen burner. Heat the crucible strongly by positioning the ring so that the bottom of the crucible is about 0.5 inches above the tip of the inner blue cone of the flame. When the magnesium ignites, place the lid, slightly ajar, on the crucible. **Warning, magnesium burns with an intensely bright flame that can damage the retinas of your eyes. Do not stare directly at the flame!**
7. When the magnesium flame subsides, remove the crucible from the flame, place it on the wire gauze and allow it to cool.
8. If necessary, delicately break the residue in the crucible into small pieces with the tip of a metal spatula. Be careful, any residue that sticks to the spatula will contribute to experimental error.
9. After the crucible is cool, use a medicine dropper or pipet to add 10 drops of distilled water to the gray residue in the crucible. Carefully smell the gas given off. Hold a piece of moist red litmus paper close to (but not touching) the mixture. Ammonia gas is basic and will turn the paper blue.
10. Place the crucible with its lid slightly ajar on the triangle over a low flame to evaporate any water present. When the sample appears dry, turn the flame to high and continue heating for five minutes.
11. Remove the crucible from the flame and allow it to cool for at least 10 minutes. Weigh the crucible and record its mass.
12. If time permits, reheat the crucible, allow it to cool and reweigh it. If the mass did not change after the second heating you have reached a constant mass. If the mass did change, repeat the process until a constant mass is achieved.
13. Dispose of the residue in the waste container provided. Clean the crucible with glass detergent (be gentle, the crucible is fragile). Rinse the crucible with distilled water, dry it with a paper towel, and return it to the instructor's cart.

Data Sheet- Experiment 4: Conservation of Mass

Name _____

Mass of empty crucible _____

Mass of crucible plus magnesium _____

Mass of Magnesium metal _____

Mass of crucible plus MgO (first heating) _____

Mass of crucible plus MgO (second heating) _____

Mass of MgO residue (Experimental Yield) _____

Calculations

Using equation 1 and the amount of Mg you used in the experiment, calculate the mass of MgO that you theoretically should have produced. This is known as the theoretical yield.

Calculate the percent yield of MgO that was actually produced according to Equation 9 below.

$$\text{Percent Yield} = (\text{Experimental Yield}/\text{Theoretical Yield}) \times 100\%$$

Discussion of Results

Write a discussion including a statement about your results and any experimental errors encountered. Discuss your results in terms of the law of conservation of mass.

Chemistry 1210 Experiment 4: Conservation of Mass, **Pre-lab**

Name _____

1. If Magnesium is burned under pure nitrogen, a black powder, Mg_3N_2 is produced. Why then is the product gray when Mg(s) is burned in air?
2. In step 8, why is it important to allow the crucible to cool before water is added? (Hint: What is likely to happen if cool water is poured into a hot porcelain crucible?)
3. Halfway through his Conservation of Mass experiment, a Joe E. Cool realized that he forgot to weigh his crucible after heating it. Short of starting over, what can Joe do to ensure accurate results? Be specific and detailed in your answer.
4. Joe burned his magnesium sample and obtained a light gray product. Since his crucible looked dirty anyway Joe skipped the step of heating to constant mass, weighed the gray sample, and calculated the mass of the product. Was his resulting mass likely to be high or low?

