EXPERIMENT 8

Analysis of a Two-Component Alloy

Objective

The goal in this experiment is to determine the mass percentage of each metal present in a sample of zinc-aluminum alloy.

Introduction

Some of the more active metals, such as magnesium, will react readily with solutions of strong acids. The products from this reaction are hydrogen gas and a solution of a salt of the metal.

$$Mg_{(s)} + 2\,H^{+}_{(aq)} \rightarrow Mg^{2+}_{(aq)} + H_{2\,(g)} \quad \textbf{(EQ 8.1)}$$

From this equation it is clear that a mole of magnesium produces one mole of hydrogen gas. If the hydrogen was collected under known conditions, it would be possible to calculate the mass of magnesium in a pure sample by measuring the amount of hydrogen gas produced by reaction with acid.

Scandium reacts spontaneously with strong acids in a manner similar to that shown by magnesium.

$$2\,Sc_{(s)} + 6\,H^{+}_{(aq)} \rightarrow 2\,Sc^{3+}_{(aq)} + 3\,H_{2\,(g)} \quad \textbf{(EQ 8.2)}$$

Using this reaction, it is possible to find the amount of scandium in a pure sample by measuring the amount of hydrogen produced by its reaction with an acid solution. In this reaction one and a half moles of hydrogen are produced by each mole of scandium that reacts.

Since the amount of hydrogen produced by a gram of Mg is not the same as the amount produced by a gram of Sc, it is possible to react an alloy of Mg and Sc of known mass with excess acid, determine the amount of hydrogen gas evolved, and calculate the percentages of Mg and Sc in the alloy using reaction equation 8.1 and 8.2 above.

In this experiment, you will react a weighed sample of a two-metal alloy with excess acid and collect the hydrogen gas evolved over water. If you measure the volume, temperature, and total pressure of gas and use the Ideal Gas Law, taking into account the pressure of water vapor in the system, you can calculate the number of moles of hydrogen gas produced by the sample.

$$PV = nRT \qquad n = \frac{PV}{RT} \tag{EQ 8.3}$$

where P is the partial pressure of hydrogen gas. The volume, V, and the temperature T of the hydrogen are easily obtained from the data. The pressure of the *dry* hydrogen, P_{H_2}, requires more attention. The total pressure P_{bar}, is, by Dalton's Law, equal to the partial pressure of the hydrogen, P_{H_2}, plus the partial pressure of the water vapor, P_{H_2O}, plus or minus the water head pressure (as explained below), $P_{\Delta h}$. For this experiment P_{bar} is equal to the measured barometric pressure.

$$P_{bar} = P_{H_2O} + P_{H_2} \pm P_{\Delta h} \tag{EQ 8.4}$$

The water vapor in the bottle is present with liquid water so the gas is saturated with water vapor; the pressure P_{H_2O} under these conditions is equal to the vapor pressure of water at the temperature of the experiment. This value is constant at a given temperature, and can be found in your textbook. Once the water vapor pressure is known it can be subtracted from equation 8.4. Now, the pressure, P_{bar} is equal to the pressure of the dry hydrogen and the head pressure.

The head pressure is measured by taking the difference in the heights of the liquid levels in the large bottle and beaker (not the reaction flask). This level represents the pressure difference in mm H_2O between atmospheric pressure, P_{bar}, and the pressure inside the large bottle and reaction flask. It is important to note that this difference must be converted to mmHg before it can be added to or subtracted from the atmospheric pressure. Since the density of Hg is 13.6 times as dense as water, the height difference values in mm must be divided by 13.6 to convert them to mmHg.

If the water level in the beaker is higher than the level in the large bottle then the pressure in the bottle is *greater* than atmospheric pressure. If the level in the beaker is below the level of the liquid in the large bottle, then the pressure in the bottle is *less than* atmospheric pressure. This will help you determine whether this difference is added to, or subtracted from the atmospheric pressure reading.

Now we know the pressure of the dry hydrogen. We now calculate the weight of each metal in the alloy as a mixture problem as the following equations illustrate:

$$\text{mass Mg} + \text{mass Sc} = \text{mass sample} = \text{mass M} \tag{EQ 8.5}$$

$$\text{mol } H_2 \text{ from Mg} + \text{mol } H_2 \text{ from Sc} = \text{mol } H_2 \text{ collected} \tag{EQ 8.6}$$

Now we have two equations and two unknowns. These can be solved simultaneously:

$$\text{mass Sc} = \text{mass M} - \text{mass Mg} \tag{EQ 8.7}$$

$$\text{mol } H_2 \text{ from Sc} = \text{mass Sc} \times \frac{1 \text{ mol Sc}}{44.96 \text{ g Sc}} \times \frac{3 \text{ mol } H_2}{2 \text{ mol Sc}} \tag{EQ 8.8}$$

$$\text{mol } H_2 \text{ from Mg} = \text{mass Mg} \times \frac{1 \text{ mol Mg}}{24.31 \text{ g Mg}} \times \frac{1 \text{ mol } H_2}{1 \text{ mol Mg}} \tag{EQ 8.9}$$

Substituting Equation 8.8 and Equation 8.9 into Equation 8.6 yields: **(EQ 8.10)**

$$\left(\text{mass Mg} \times \frac{1\ \text{mol Mg}}{24.31\ \text{g Mg}} \times \frac{1\ \text{mol H}_2}{1\ \text{mol Mg}}\right) + \left(\text{mass Sc} \times \frac{1\ \text{mol Sc}}{44.96\ \text{g Sc}} \times \frac{3\ \text{mol H}_2}{2\ \text{mol Sc}}\right) = \frac{PV}{RT} \quad \textbf{(EQ 8.11)}$$

Using Equation 8.7 we can substitute for the mass of Sc into Equation 8.11:

$$\left(\text{mass Mg} \times \frac{1\ \text{mol Mg}}{24.31\ \text{g Mg}} \times \frac{1\ \text{mol H}_2}{1\ \text{mol Mg}}\right) + \left((\text{mass M -mass Mg}) \times \frac{1\ \text{mol Sc}}{44.96\ \text{g Sc}} \times \frac{3\ \text{mol H}_2}{2\ \text{mol Sc}}\right) = \frac{PV}{RT} \quad \textbf{(EQ 8.12)}$$

Solving for the mass of Mg in Equation 8.12 gives the grams of metal, *M*, in the alloy sample. Substituting this into Equation 8.7 gives the grams of Sc in the original sample. The percentages of each of the metals can be calculated in the following manner.

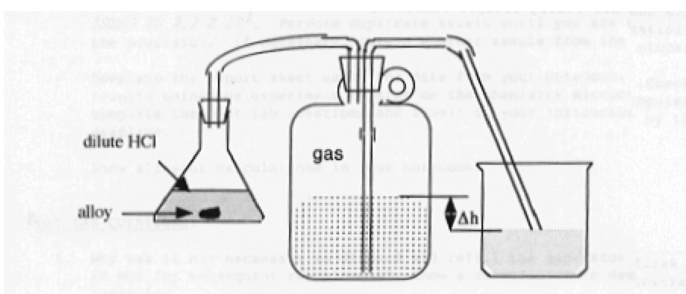
$$\frac{\text{mass Mg}}{\text{mass sample}} \times 100 = \%\text{Mg} \quad \textbf{(EQ 8.13)}$$

$$\frac{\text{mass Sc}}{\text{mass sample}} \times 100 = \%\text{Sc} \quad \textbf{(EQ 8.14)}$$

Procedure

1. Obtain an unknown Zn-Al alloy sample and requisite apparatus from the stockroom. Assemble the apparatus illustrated in Figure 8.1 :

FIGURE 8.1



2. A demonstration apparatus should also be displayed on your instructor's bench. The generator flask is a 250 mL Erlenmeyer flask fitted with a one-hole stopper; the beaker is a 1000 mL beaker. Insert the glass tubes in the two-hole stopper of the large bottle with equal treatment and care. The generator flask should be placed in a beaker.

> **CAUTION:** Remember to lubricate the hole of the stopper with glycerin prior to inserting the glass tube.

3. Determine the mass of a piece of weighing paper to the maximum precision of the electronic balance (this is called the “tare”). Do the same for a second piece of weighing paper and place one piece of your unknown on the first paper and another piece on the second paper. Weigh each again and record these values.
4. Fill the large bottle to near the top with water. Determine the mass of the bottle and water on a balance of adequate capacity. Assemble the apparatus for the first experimental run.

5. Add 50-75 mL of dilute (6 *M*) HCl to the generator flask.
6. Add your first alloy sample to the generator flask and immediately insert the stopper. The gas will flow through the tubing and displace the water from the large bottle. Swirl the flask as needed to continue the reaction until all of the alloy is reacted. Allow the system to equilibrate for 10 minutes after the reaction is complete.
7. Measure the difference in mm of the water level in the bottle and beaker. Remove the stopper from the large bottle and measure the temperature of the remaining water and reweigh it on the balance. The difference between the initial and final mass of the bottle and water is the mass of water displaced by the gas. Using the density for water (look up the value in the CRC handbook of Chemistry and Physics), the volume of the gas displaced can be determined.
8. Take the temperature of the water and the temperature of the gas above the water.
9. Repeat the experiment with the second sample of alloy using the acid solution from the previous trial. When you are completely finished with the experiment, dump the bottle full of water in the planter outside.

NOTE: For an approximate quick check of the precision of your results, determine the ratio of the mL of H_2 produced to the grams of alloy used for each trial. Excellent precision should give ratio values that are equal to two significant figures. For example, the values, 818 and 823, both round to 8.2×10^2. Perform duplicate trials until you are satisfied with the precision of your results.

Results and Calculations

In the calculations show:

1. how you derived the equation to determine the amount of zinc in the sample from the mass of the alloy and the total number of moles of hydrogen gas from the reaction. Do this without putting the **actual mass** of alloy or **number of moles** of gas from the data;
2. a sample calculation using this equation on one of your trials;
3. create a table summarizing your major data.

Post Lab Questions

1. Why was it not necessary to discard and refill the generator flask with fresh 6 *M* HCl for subsequent reactions? Show a calculation to demonstrate your reasoning.

2. In an experiment similar to yours, students were given an unknown sample that produced H_2S or NH_3 gas. What problem(s), if any, might arise if the same procedures as this experiment were followed?

3. Why is it not necessary to consider the amount of air that was in the generator flask at the start of the reaction?

4. If there was a leak in the rubber tubing that allowed approximately 150 mL of H_2 to escape, how would it affect the relative amounts (% values) that you calculated as the results? Would the %zinc in the sample appear to be high, low, or remain the same? Show a calculation to justify your answer.

5. What assumptions do we make regarding the temperatures of the generator and collection flasks? Are they valid?

6. How would the calculated number of moles of H_2 collected change if the generator flask were not given time to fully equilibrate?