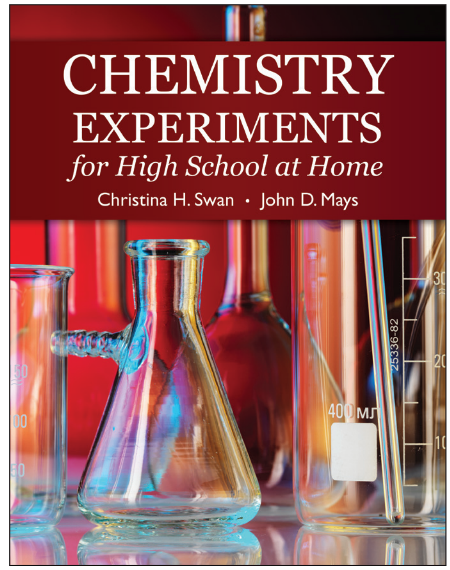


***Teacher Guide***

***for***

## Chemistry Experiments for High School at Home

***by***

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## Introductory Note

This document is a work in progress. We are adding to it as questions arise. Please contact science@classicalsubjects.com with any questions you have while working through the chemistry experiments. We are happy to assist!

In this document, the following acronyms are used:

CEHSH *Chemistry Experiments for High School at Home*

GC *General Chemistry* (also applies to *Principles of Chemistry*, published by Centripetal Press)

CAS *Chemistry for Accelerated Students* (also applies to *Accelerated Chemistry*, published by Centripetal Press)

## Experiment 9

### *Pre-Lab Calculations*

In this pre-lab, we need not worry too much about significant digits in the calculations since we are dealing only with approximate quantities anyway. First, we must balance the given formula equation so that we have the correct mole ratios for the reaction. The balanced equation is:

$$2\text{Al}(s) + 6\text{HCl}(aq) \rightarrow 2\text{AlCl}_3(aq) + 3\text{H}_2(g)$$

Next, let's calculate the approximate mole quantities of the reactants, as they are described in the Part 1 procedure. For the 6.0 *M* HCl:

$$10 \text{ mL HCl} \cdot \frac{1 \text{ L}}{1000 \text{ mL}} \cdot \frac{6.0 \text{ mol}}{\text{L}} = 0.06 \text{ mol HCl}$$

For the 0.25 g of aluminum:

$$0.25 \text{ g Al} \cdot \frac{1 \text{ mol}}{26.98 \text{ g}} = 0.0093 \text{ mol Al}$$

From the balanced equation, the mole ratio of Al : HCl is 1 : 3. To consume 0.0093 mol Al we need 0.0093 · 3 = 0.028 mol HCl. We have about twice this amount, so Al is the limiting reactant. (Note: This excess acid is why the cleanup procedure for Part 1 says the beaker contains concentrated hydrochloric acid—all the aluminum is consumed in the reaction, but not all the HCl is consumed.)

To predict the quantity of $H_2$ produced by the reaction, we note from the balanced equation that the mole ratio of Al to $H_2$ is 2 : 3. Thus, for every two moles of Al going into the reaction, 3 mol $H_2$ are produced. In other words:

$$\# \text{ moles } H_2 = \frac{3}{2} \cdot \# \text{ moles Al}$$

Thus, the predicted amount of $H_2$ produced by this reaction is:

$$\# \text{ moles } H_2 = 1.5 \cdot 0.0093 = 0.014 \text{ mol } H_2$$

Now we must calculate the volume of $H_2$ by using the law of partial pressures, estimated lab conditions, and the ideal gas law.

The law of partial pressures applied here gives us:

$$P_T = P_{gas} + P_{water\ vapor}$$

From this, the pressure of the hydrogen gas produced is

$$P_{gas} = P_T - P_{water\ vapor}$$

Assuming a temperature of approximately 22°C, Table 9.1 in CEHSH (and Table A.4 in GC or B.4 in CAS) indicates that the vapor pressure of water is about 20 Torr. Assuming a barometric pressure in the lab of 760 Torr, we have

$$P_{gas} = 760 \text{ Torr} - 20 \text{ Torr} = 740 \text{ Torr}$$

and

$$T_{gas} = 22°\text{C} + 273.15 = 295.15 \text{ K}$$

Now we can use the ideal gas law to calculate the approximate volume of the $H_2$ produced by the reaction.

$$PV = nRT$$

$$V = \frac{nRT}{P} = \frac{0.014 \text{ mol} \cdot 62.4 \frac{\text{L} \cdot \text{Torr}}{\text{mol} \cdot \text{K}} \cdot 295.15 \text{ K}}{740 \text{ Torr}} = 0.348 \text{ L} = 348 \text{ mL}$$

### *Part 2 Calculations*

In this section, we go through the calculations for the acetic acid and baking soda reaction. Students should perform these calculations using the precise quantities measured during the experiment. We do the calculations here using the approximate quantities given in the procedure.

The balanced equation is the same as the formula equation:

$$CH_3COOH(aq) + NaHCO_3(s) \rightarrow NaCH_3COO(aq) + H_2O(l) + CO_2(g)$$

From this equation, we see that all the mole ratios are 1 : 1.

In the early printings of CEHSH, the acetic acid concentration was listed as 0.1 *M*. However, this was changed to 1.0 *M* in later printings. Using this revised concentration, the reactant mole quantities are calculated as follows:

$$30\ \text{mL}\ CH_3COOH \cdot \frac{1\ \text{L}}{1000\ \text{mL}} \cdot \frac{1\ \text{mol}}{\text{L}} = 0.030\ \text{mol}\ CH_3COOH$$

$$1\ \text{g}\ NaHCO_3 \cdot \frac{1\ \text{mol}}{84\ \text{g}} = 0.0119\ \text{mol}\ NaHCO_3$$

Knowing that the mole ratios are 1 : 1, these reactant quantities indicate that $NaHCO_3$ is the limiting reactant and that we expect 0.0119 mol $CO_2$ to be produced by the reaction. For our calculation we will use the same temperature and pressure as before. Again, students should use data collected in the experiment.

$$PV = nRT$$

$$V = \frac{nRT}{P} = \frac{0.0119\ \text{mol} \cdot 62.4\ \frac{\text{L} \cdot \text{Torr}}{\text{mol} \cdot \text{K}} \cdot 295.15\ \text{K}}{740\ \text{Torr}} = 0.296\ \text{L} = 296\ \text{mL}$$