Measurements Using Electrochemical Cells and Electroplating

The basic counting unit in chemistry, the mole, has a special name, *Avogadro's number*, in honor of the Italian scientist Amadeo Avogadro (1776-1856). The commonly accepted definition of Avogadro's number is the number of atoms in exactly 12 g of the isotope ^{12}C, and the quantity itself is $6.02214199(47) \times 10^{23}$.[1]

A bit of information about Avogadro seems appropriate. His full name was Lorenzo Romano Amedeo Carlo Avogadro (almost a mole of letters in his name). He was a practicing lawyer until 1806 when he began his new career teaching physics and math at the University of Turin, where he was later promoted to the chair of physical chemistry. In 1811, Avogadro published a paper in the *Journal de Physique*, entitled "Essay on a Manner of Determining the Relative Masses of the Elementary Molecules of Bodies, and the Proportions in Which They Enter into These Compounds," which pretty much says it all. This paper includes the statement that has come to be regarded as Avogadro's Hypothesis:

> The first hypothesis to present itself in this connection, and apparently even the only admissible one, is the supposition that the number of integral molecules in any gases is always the same for equal volumes, or always proportional to the volumes. Indeed, if we were to suppose that the number of molecules contained in a given volume were different for different gases, it would scarcely be possible to conceive that the law regulating the distance of molecules could give in all cases relations as simple as those which the facts just detailed compel us to acknowledge between the volumes and the number of molecules.

In this experiment, you will confirm Avogadro's number by conducting an electrochemical process called *electrolysis*. In electrolysis, an external power supply is used to drive an otherwise *nonspontaneous* reaction. You will use a copper strip and a zinc strip as the electrodes, placed in a beaker of sulfuric acid. You will make the cell electrolytic by using the copper strip as the anode and the zinc strip as the cathode. By determining the average current used in the reaction, along with the knowledge that all of the copper ions formed are the 2^+ cations, you will calculate the number of atoms in one mole of copper and compare this value with Avogadro's number.
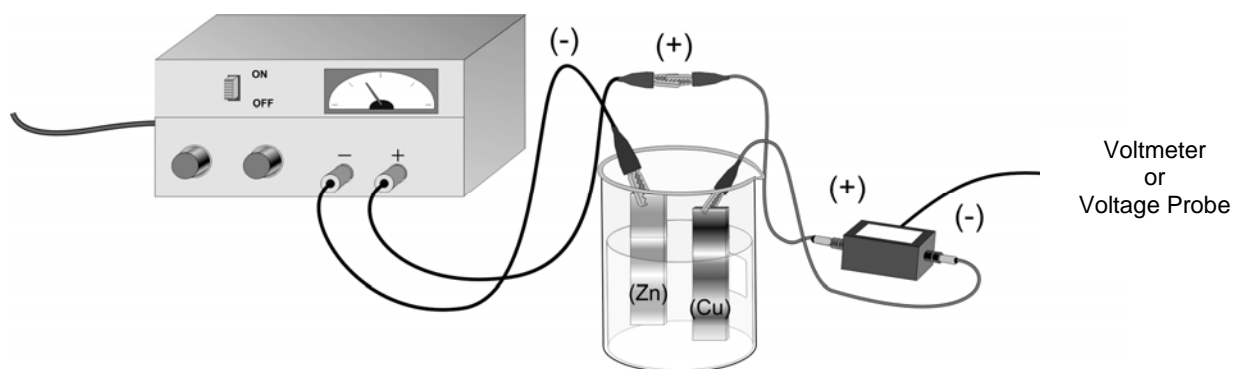

Figure 1

OBJECTIVES

In this experiment, you will

- Prepare an electrochemical cell to oxidize a copper electrode.
- Measure the amount of copper that was deposited in the electroplating process and determine the average current used.
- Calculate a value for Avogadro's number and compare it to the accepted value.
- Calculate the value of a Faraday.

MATERIALS

Ammeter or Data Collection Mechanism with computer or handheld	1 M H_2SO_4 solution
Vernier Current Probe if not using voltmeter	copper strip (anode) (+)*
1.5 volt DC power source (check with physics or AP Biology, can also use 2 D cells in series)	zinc strip (cathode) (−)*
	distilled water
3 connecting wires with alligator clips (check with physics)	250 mL beaker
steel wool or sandpaper	analytical balance

* Remember the acronym "**EPA**"—An **E**lectrolytic (nonspontaneous) cell has a **P**ositive **A**node.

PROCEDURE

1. Obtain and wear goggles.

2. Use steel wool or sandpaper to clean a strip of copper, which will be the anode of the electrochemical cell. Obtain a strip of zinc metal to use as the cathode, and clean it with steel wool if needed.

3. Use an analytical balance to measure the mass of the copper strip. Record the mass in your data table.

4. Fill a 250 mL beaker about ¾ full with 1 M H_2SO_4 solution. **CAUTION:** *Handle the sulfuric acid with care. It can cause painful burns if it comes in contact with the skin.*

5. Obtain a DC power supply and either an Ammeter or Current Probe. Use connecting wires, with alligator clips, to connect the DC power supply, ammeter or current probe, and the two metal electrodes to be used in the electrochemical cell as you see in *Figure 1*. Copper is the anode (+) terminal and zinc is the cathode (−) terminal in this cell because it is an electrolytic cell and the polarities of the electrodes are switched since it is a *nonspontaneous* process. **Note: Do not place the electrodes in the cell until Step 7.**

6. Set up the ammeter or data collection system. If using a data collection system, follow these steps:

   a. Connect the Current Sensor to the interface.

   b. Start the data collection program.

   c. Set up data collection for a Time Graph, taking a sample every 5 seconds and 36 samples. Data will be gathered for 3 minutes. If using an ammeter, manually collect data for 3 minutes.

7. Place the electrodes into the 1 *M* H_2SO_4 solution in the cell. Make sure that the electrodes are immersed in the solution to equal depths and as far apart as possible.

8. Turn on the DC power supply and check the current readings. The initial current should be in the 0.4 – 0.6 amp range. If the current is not in this range, adjust the settings on the power supply. Once the initial current is in range, turn off the power supply.

9. Begin the data collection and turn on the power source. Data will be collected for 3 minutes. Observe the reaction carefully. **Note:** Be ready to shut off the power as soon as the data collection stops.

10. When the data collection is complete, turn off the power supply and carefully remove the electrodes from the H_2SO_4 solution. Carefully rinse the copper electrode with distilled water. Dry the copper electrode very carefully.

11. Measure and record the mass of the *dry* copper electrode.

12. Analyze the graph of your data to determine the average current that was applied during the electrolysis. Record the mean in your data table as the average current for Trial 1.

13. Repeat the necessary steps to conduct a second trial. At the direction of your instructor, conduct a third trial.

DATA TABLE

	Trial 1	Trial 2	Trial 3 if time allows
Initial mass of copper electrode (g)			
Final mass of copper electrode (g)			
Average current (A)			
Time of current application (s)			

PRE-LAB QUESTIONS

1. What is the sign of $E°_{cell}$ for the reaction taking place in this laboratory exercise? Justify your answer.

2. What is the sign of $\Delta G°$ for the reaction taking place in this laboratory exercise? Mathematically justify your answer.

3. In this experiment, which electrode is losing mass? Identify whether this electrode is the anode or the cathode. Justify your answer by writing the half-reaction that occurs at the electrode.

4. Answer the following questions about electrolytic cells.

    (a) Which redox process occurs at the anode?

    (b) What is the polarity of the anode?

    (c) Which direction do electrons flow?

    (d) In which direction do cations migrate?

5. A student conducted an experiment to determine Avogadro's number using a zinc cathode and copper anode having an initial mass of 5.100 g. The student applied a current of 0.468 amps for a total of three minutes. After data collection was complete, the copper anode was dried and massed. The final mass of the copper anode was 5.073 g.

    (a) Calculate the number of grams lost by the copper anode.

    (b) Calculate the total quantity of charge in coulombs that passed through the electrolytic cell.

    (c) Calculate the number of electrons used to electroplate the zinc. The charge of a single electron is 1.602×10^{-19} coulombs.

    (d) Calculate the number of copper atoms lost from the copper electrode.

    (e) Calculate the number of copper atoms per gram of copper lost at the anode.

    (f) Calculate the number of copper atoms in a mole of copper.

    (g) Calculate the student's percent error.

    (h) Starting with your answer to part (b) above, show one dimensional analysis expression that allows you to arrive at your answer to part (f).

6. A few seconds pass between the end of data collection and the time the power supply is turned off. What effect does this error have on the calculated value of Avogadro's number?

POST-LAB QUESTIONS AND DATA ANALYSIS

1. Calculate Avogadro's number for each trial.

2. Calculate the percent error for each trial

3. Which measurement limited the number of significant figures that could be obtained?

5. The Faraday is a unit of measure that relates the number of coulombs per mole of electrons. Use your data to calculate the value of the Faraday for each trial. The accepted value of the Faraday is 96,500 C/ mol e^-.

7. A student fails to dry the electrode completely after Trial 1prior to massing it on the balance.

   (a) What effect does this error have on the calculated value of the Faraday?

   (b) What effect does this error have on the calculated value for Avogadro's number for Trial 1?