Experiment 18
Experimental Determination of Avogadro's Number

Objectives:

- To construct an electrochemical cell (also referred to as battery).
- To measure the current produced in this cell.
- To use stoichiometry to determine the theoretical change in mass of electrode.
- To use the previous objectives to determine Avogadro's number and the percentage difference from the accepted value.

To prepare for this experiment:

- Carefully read the entire experimental guide (below).
- Answer all the prelaboratory practice problems.
- Since some of this material is not covered in lecture, a short introductory lesson providing the necessary fundamental concepts will be provided. So listen closely!
- Review dimensional analysis.

Introduction:

An electrochemical cell is a device that uses chemical reactions to convert chemical energy into electrical energy. These types of devices are seen often in the form of batteries, which power everyday electronics, such as laptops and cell phones. In this lab, you will be creating a variation of what is known as a Galvanic Cell. In this type of electrochemical cell, two different metal electrodes are submerged in a solution that contains the aqueous ionic form of the same two metals. This is demonstrated below in Figure 1.

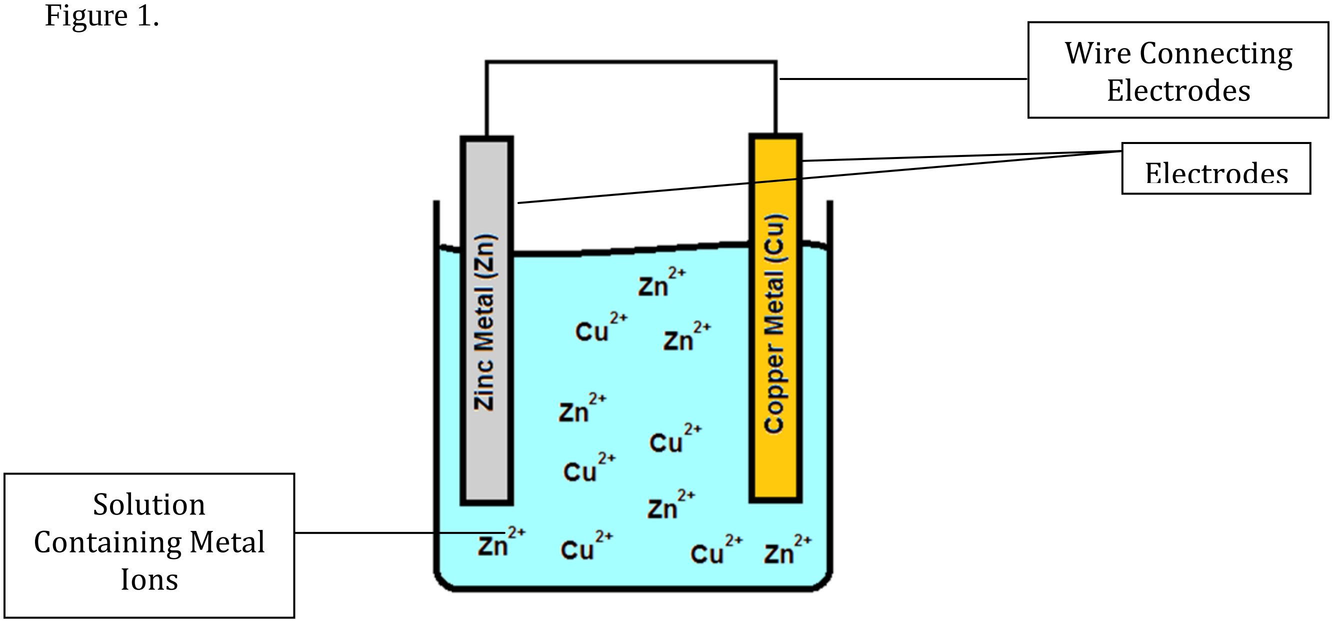

Figure 1: Electrochemical Cell

Electrochemical cells utilize Redox (Reductions-Oxidation) reactions. During reduction, an aqueous ion accepts electrons and becomes neutral, thus becoming solid metal. In oxidation, electrons are liberated by solid metals, thus turning them into ions that become dissolved in solution. In this experiment, **Copper is Reduced** and **Zinc is Oxidized**. So as the reactions proceed, $Cu_{(s)}$ is being formed and $Zn_{(s)}$ is being consumed.

The wire shown in Figure 1 attaches the electrodes to allow for the flow of electrons, which is known as a current. The unit of current is Amperes (often called Amps for short). This unit measures the charge of the electrons over time, where charge is measured in Coulombs. Using this, you can obtain the following equation to relate charge to time:

Charge = Current * Time

where charge is measured in Coulombs (C), current is measured in Amps (A = Coulomb/Second), and time is measured in seconds. As you can see, this is consistent with the dimensional analysis.

$$C = A * s = \left(\frac{C}{s}\right) * s$$

It may not be obvious, but all of this information can be used to determine Avogadro's number, which gives the number of molecules or atoms in a mole of any substance. All electrons have a constant charge of $1.602*10^{-19}$ C. If you measure the current of this system for a certain amount of time, you can determine the total charge passed through the wire, thus determine the total number of electrons that were involved in the reaction.

$$\frac{Total\ Charge\ Passed}{Charge\ of\ a\ Single\ Electron} = Number\ of\ Electrons\ Involved$$

This number of electrons can be correlated to the theoretical number of Zinc or Copper atoms involved using stoichiometry. **For this experiment, we will choose to correlate the number of electrons involved to the number of Copper atoms involved.** Since Copper metal is deposited, you can measure the change in mass of the Copper electrode, which can be converted to moles. The ratio of atoms of Copper to moles of Copper should tell you how many atoms are in a mole of Copper, which is Avogadro's number.

$$\frac{Atoms\ of\ Cu}{Moles\ of\ Cu} = Avogadro's\ Number$$

Procedures

Part A: Setting up the Galvanic Cell

1. Fill a 250 mL beaker with 175 mL of 0.5 M copper sulfate solution.
2. Use the steel wool to carefully clean the zinc and copper metal strips. These will serve as electrodes (zinc strip is the anode and copper strip is the cathode).
3. Measure the mass of the copper strip to the nearest 0.0001 g using an analytical balance. Record the mass in your data table.

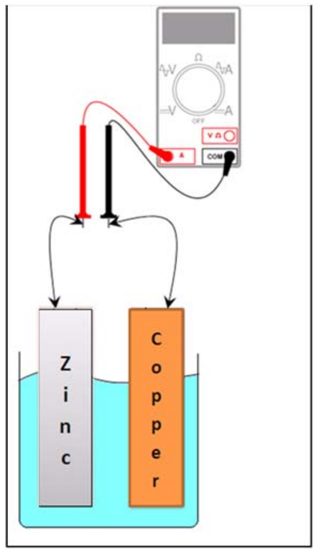

Figure 2: Galvanic Cell setup with Ammeter

4. Attach one end of an alligator clip to the top of the copper electrode. Attach the other end to top of the zinc electrode, as shown in Figure 2.
5. Connect the electrodes to the ammeter probes using 2 alligator clips as shown in figure 2.

Part B: Data Collection

1. Turn on ammeter and set it to 200 mA.
2. Use a stopwatch to keep time.
3. Have one student place the electrodes into the solution, making sure the two electrodes do **not** touch each other. The other student should begin timing and recording current as soon as the electrodes are in the solution
4. Record the current reading (in mA) every 15 seconds in Data Table 2.
5. After 5 minutes, carefully remove the electrodes from the solution.
6. Carefully pat the copper electrode dry using a paper towel.
7. Measure the mass of the dried copper electrode to the nearest 0.0001 g and record in Data Table 1.
8. Clean the zinc electrode with a paper towel.
9. Repeat steps 2-4 in Part A and steps 2-7 in Part B.

Part C: Cleaning Up

1. Dispose of the copper sulfate solution in the inorganic waste container.
2. Clean both the copper and zinc electrodes.
3. Dispose of all paper towels in the garbage.

Data Table 1

	Trial 1	Trial 2
Initial Mass of Copper Electrode (g)		
Final Mass of Copper Electrode (g)		
Average Current (mA)		
Total Time (sec)	**300**	**300**

Data Table 2

Time (s)	**Trial 1 Current (mA)**	**Trial 2 Current (mA)**
0		
15		
30		
45		
60		
75		
90		

Time (s)	**Trial 1 Current (mA)**	**Trial 2 Current (mA)**
105		
120		
135		
150		
165		
180		
195		

Time (s)	**Trial 1 Current (mA)**	**Trial 2 Current (mA)**
210		
225		
240		
255		
270		
285		
300		

Average Current		

<u>Data Analysis:</u>

1. Use the average value for your current measurements (in Amps) over the course of the reaction to find the total amount of charge (in Coulombs) passed through the circuit.
2. Calculate the number of electrons involved in the reaction. (Reminder: The charge of an electron is $1.602*10^{-19}$C)
3. Using the answer from #2, determine the number of Copper atoms deposited. This will be based on the stoichiometry of the reaction. (2 electrons are used to reduce 1 atom of Cu^{2+})

4. Find the difference in the final and initial mass of the electrode, and use this to determine the number of moles gained from the electrolysis reaction. (For the molar mass of Cu use 63.546)
5. Using the ratio described in the Introduction and previous calculations, determine your experimental Avogadro's Number
6. Calculate percent error from the theoretical value of $6.02*10^{23}$. For your percent error calculation you may use the formula given below:

$$Percent\ Error = \frac{experimental\ value - theoretical\ value}{theoretical\ value}\ x\ 100$$

7. Repeat these calculations (#1-#6) for your second trial.

Pre-lab Questions

1. Galvanic cells convert __________ energy to __________ energy.

   a) chemical; electrical
   b) thermal; electrical
   c) chemical; mechanical
   d) ionic; mechanical

2. What are the two metals that will be used as electrodes?

   a) Copper and iron
   b) Nickel and iron
   c) Copper and zinc
   d) Nickel and zinc

3. The “redox” reaction gets its name because it involves:

   a) A red colored bovine
   b) Reduction and oxidation
   c) A red colored oxide compound
   d) Rhenium, deuterium, and oxygen

4. An ampere is a unit of:

   a) Charge
   b) Time
   c) Energy
   d) Current

5. Avogadro’s number is the number of __________ in a __________.

   a) atoms; mole
   b) grams; mole
   c) atoms; gram
   d) moles; gram

Post-lab Questions

1. Use the steps in the data analysis section to determine the following;
    a. Determine the moles of copper gained the copper electrode in each trial.

    b. Determine the amount electrons passed through the multimeter for each trial.

    c. Determine the number of copper atoms theoretically involved in each trial

    d. Determine Avogadro's number for each trial.

2. Calculate the percent error for each trial, also using the steps outlined in the data analysis section. Use 6.022×10^{23} as the theoretical value for Avogadro's number.

3. What are some sources of error in this lab?

4. Could you have determined Avogadro's number using the mass change in the Zinc Electrode instead? If so, briefly explain how.