Determining Avogadro's Number

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. The coulomb (symbolized C) is the standard unit of electric charge in the International System of Units (SI). In terms of SI base units, the coulomb is the equivalent of one ampere-second.

$$\mathbf{C = As}$$

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.

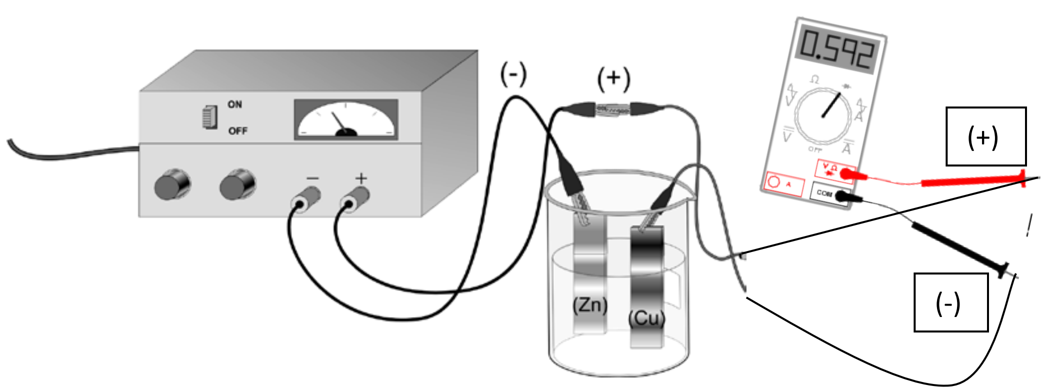

Figure 1

[1] *CRC Handbook of Chemistry and Physics*, 82nd edition, pp. 1-7.

OBJECTIVES

In this experiment, you will, in groups of 3 to 4 students, find the conditions which gives optimal value for Avogadro's number and then write a brief conclusion based on your findings. A minimum of 6 but maximum of 9 trials should be performed; changing ONLY ONE parameter at a time.

- Prepare an electrochemical cell to **oxidize** a copper electrode. ONCE THE CIRCUIT IS COMPLETE, BEFORE STARTING, CHECK IF THE VOLTMETER IS REGISTERING A VALUE BETWEEN 0.2 – 0.6 A. If not, then a wire might be damaged, so contact your instructor.
- 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 under different conditions then use that knowledge to determine optimal set-up conditions. Experimental variables which may or may not alter the results:
  1) Does the color of the wire make any difference in the value obtained?
  2) 0.2 -0.6 Amps is requested; which value works best? 0.2 or 0.6?
  3) Will the amount of gas produced be a depended on the amperage? 0.2 or 0.6?
  4) Will the spacing between electrodes effect the value?
  5) Will cleaning electrodes at different points throughout the experiment vary the results?
  6) Will wiping the copper electrode hard or softly vary the results?
  7) Will turning the Amps on before dipping the electrodes compared to after dipping, change the results? (Ask yourself, does order of set-up matter?)
  8) If a solid forms in the beaker, what does that represent and how will it affect the value?
  9) Are all of the wires connected correctly? Will making small adjustments to how the wires and alligator clips are attached vary the results?
  10) What color was the solution by the end of the experiment? What role does the color of the solution represent?

MATERIALS

Voltmeter
1.5 volt DC power source
6 connecting wires with alligator clips
steel wool
analytical balance
1 M sulfuric acid, H_2SO_4, solution
copper strip (anode)
zinc strip (cathode)
distilled water
two 250 mL beakers

BASIC PROCEDURE

1. Obtain and wear goggles.

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

3. Use an analytical balance to measure the mass of ONE of the copper strips. Record the mass in your data table. The other strip is for your initial set-up test.

4. Fill a 150 or 250 mL beaker about $^1/_2$ 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 a Voltmeter from the stockroom. Use the connecting alligator clip wires to connect the DC power supply to the Voltmeter, and to the two metal electrodes used in the electrochemical cell as shown in Figure 1. DO NOT ALLOW THE ALLIGATOR CLIPS TO HAVE CONTACT WITH ANY SOLUTION/LIQUIDS! Never submerge the alligator clips. Copper is the anode and zinc is the cathode in this cell because we do not want the reaction to occur spontaneously. **Note: Do not place the electrodes in the cell for data collection until Step 7.**

6a. **FIRST** test to see if your set-up works. Place the electrodes (not the one that you already weighed) 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.

6b. Turn on the DC power supply and check the current readings. The initial current should be in the 0.2–0.6 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.

7. Using a stopwatch – collect data for about 180 sec (3 min), record the exact time; recording the amps approximately every 15-20 sec. Connect the weighed copper electrode. Start data collection by turning on the power source. Data will be collected for 3 minutes. Observe the reaction carefully. **Note:** Be ready to turn off the power as soon as the data collection stops.

8. 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.

9. Measure and record the mass of the dry copper electrode.

10. Determine the average current applied during the experiment.

11. Reconnect the electrodes and repeat Steps 7–10 to conduct another trial. Continue making single changes as described in the OBJECTIVES section and record the effect each change.

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REPORT SHEET

Name: ______________________________

Section: ____________________

DATA TABLE OF MEASUREMENTS

	Trial 1	Trial 2	Trial 3	Trial 4	Trial 5	Trial 6	Trial 7	Trial 8
Initial mass of copper electrode (g)								
Final mass of copper electrode (g)								
Average current (A)								
Time of current application (s)								

All data & calculations should be recorded in your lab notebook, so attach the carbon copy to this report sheet. Down below only show a sample calculation for one trial.

DATA ANALYSIS

1. Calculate the total charge, in coulombs (C), which passed through the electrolytic cell for each trial. (C = As)

2. Use your answers from question 1 above to calculate the number of electrons in the electrolysis for each trial. Recall from the famous Millikan oil-drop experiment that the charge of an electron is 1.602×10^{-19} C/e (Coulombs per electron).

3. Determine the number of copper atoms lost from the anode in each trial. Remember that the electrolysis process uses two electrons to produce one copper ion (Cu^{2+}).

4. Calculate the number of copper atoms per gram of copper lost at the anode for each trial (#atoms/g). The mass lost at the anode is equal to both the mass of copper atoms lost and the mass of copper ions produced (the mass of the electrons is negligible).

5. Calculate the number of copper atoms in a mole of copper for each trial by using the molar mass. [atoms/g * g/mol = atoms/mol]   Compare this experimental value to Avogadro's number.

6. Write a conclusion based on your findings, describing the optimal conditions needed for this experiment and anything else discovered during this experiment. Include the names your fellow collaborators.