

DETERMINING AVOGADRO'S NUMBER BY ELECTROLYSIS

LAB ADV COMP 31

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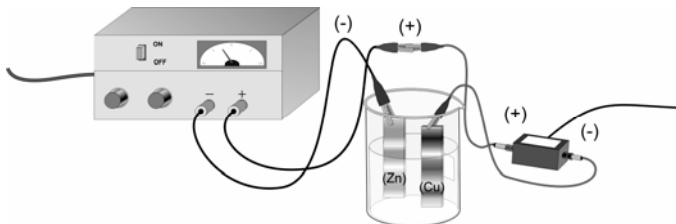
INTRODUCTION

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.



¹ *CRC Handbook of Chemistry and Physics*, 82nd edition, p. 1-7.

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.

MATERIALS

Vernier computer interface	1 M H ₂ SO ₄ solution
computer	copper strip (anode)
Vernier Current Probe	zinc strip (cathode)
1.5 volt DC power source	distilled water
4 connecting wires with alligator clips	two 250 mL beakers
steel wool	analytical balance

PROCEDURE

1. Obtain and wear goggles.
2. Use steel wool 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 $\frac{3}{4}$ full with 1 M H₂SO₄ 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 Vernier Current Probe. Use connecting wires, with alligator clips, to connect the DC power supply, Current Probe, and the two metal electrodes to be used in the electrochemical cell (see Figure 1). Copper is the anode and zinc is the cathode in this cell because you do not want the reaction to occur spontaneously. **Note:** Do not place the electrodes in the cell until Step 8.
6. Connect the Current Probe to Channel 1 of the Vernier computer interface. Connect the interface to your computer with the proper cable.
7. Start the *Logger Pro* program on your computer. Open the file "31 Avogadro" from the *Advanced Chemistry with Vernier* folder.
8. Place the electrodes into the 1 M H₂SO₄ solution in the cell. Make sure that the electrodes are immersed in the solution to equal depths and as far apart as possible.
9. Turn on the DC power supply and check the current readings. The initial current should be in the 0.2 – 0.6 A 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.

10. Click  Collect to begin the data collection and turn on the power supply. 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.
11. When the data collection is complete, turn off the power supply and carefully remove the electrodes from the H₂SO₄ solution. Carefully rinse the copper electrode with distilled water. Dry the copper electrode very carefully.
12. Measure and record the mass of the dry copper electrode, in your data table, for Trial 1.
13. Analyze the graph of your data to determine the average current that was applied during the electrolysis. Click the Statistics button, . Record the mean in your data table as the average current for Trial 1.
14. Repeat Steps 8-13 to conduct a second trial. At the direction of your instructor, conduct a third trial.

DATA TABLE

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

DATA ANALYSIS

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

2. Use your answers from 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} coulombs per electron.

Determining Avogadro's Number

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. 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. Compare this value to Avogadro's number.