

Experiment 11

The Determination of Avogadro's Number

Background

It is hard to imagine the enormity of Avogadro's number. Did you know that Avogadro's number of water droplets would cover all the land in the United States to a depth of two miles? Avogadro's number of pennies placed in a rectangular stack 6 meters x 6 meters at the base would stretch 1×10^{13} km high and extend outside our solar system. It would take light nearly a year to travel from one end of the stack to the other.¹

Amadeo Avogadro (1776-1856) never knew the number 6.022×10^{23} . It is important to note that Avogadro did not invent the mole. (The term "mole" was not coined until the early 1900s.) However, Avogadro's hypothesis helped chemists determine a system of relative molecular weights for different gases. For example, imagine you had a box of apples that weighed twice the amount of a box of oranges. Could you say that an apple is twice as heavy as an orange? You could only say that if you knew there was the same number of apples as oranges in the boxes. Avogadro allowed chemists to use volume to gauge the relative number of particles in order to allow a mass comparison.

In chemistry, the mole is an important unit of measure. In Latin, *mole* means a "massive heap", which is how it should be thought of in chemistry—as a massive heap of very tiny, invisible particles, quantified so they become visible. The mole simply refers to an amount of a substance. Specifically, a mole of any substance is its relative atomic mass (or formula mass) in grams. Its precise definition is this:

The mole is the amount of a substance of a system which contains as many elementary entities as there are atoms in 0.012 kilogram of carbon-12; its symbol is "mol." When the mole is used, the elementary entities must be specified and may be atoms, molecules, ions, electrons, other particles, or specified groups of such particles.²

Avogadro's number³, N_A represents the number of particles in a mole and has been determined by many different experimentalists, notably Jean Perrin (1870-1942) and Robert Millikan (1868-1953). It was even calculated by Albert Einstein (1879-1955) in his Ph.D. dissertation! The most accurate number to date represents the number of particles in one mole: $N_A = 6.02214199 \times 10^{23}$ and is achieved by X-ray crystallography data.

This Experiment

Unlike most units of measure with which we are familiar (like the dozen), a mole of particles cannot be counted—it's too large! It is determined experimentally by indirect means, in this case, by an **electrolysis** experiment. To get a very precise value of Avogadro's number, we would need some expensive equipment. However, there are several ways that students can get a reasonably accurate value by inexpensive means, including the experiment described in the following pages.

Safety Precautions

Sulfuric acid is corrosive. In case of contact with skin, rinse with plenty of water and notify your instructor. Wear goggles at all times in the chemistry laboratory.

¹ <http://www.chemistryexplained.com/Ma-Na/Mole-Concept.html>. Accessed 8/27/2006.

² National Institute of Standards and Technology (NIST), 14th General Conference, 1971. <http://physics.nist.gov/cuu/Units/mole.html>. Accessed 8/27/2006.

³ This number is also referred to as Loschmidt's number, L , named after German chemist Josef Loschmidt who published an article in 1865 that led to the determination of the number of molecules/cm³ of a gas.

Experiment 3 Turn-in sheets

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Experimental Procedure

Note: Read all directions carefully. Ask your instructor if you need assistance.

Caution: Sulfuric acid is corrosive and can cause burns. Handle with care. In case of contact with skin, rinse thoroughly. You must wear goggles when handling acid.

Setup the electrolysis:

1. Obtain and wear goggles! Use a ring clamp, ring stand, masking tape, alligator clips and 2 copper electrodes to set up your electrolysis. Lower the electrodes into a 150-mL beaker. Make sure the power supply is off and unplugged during this time. Do not allow the electrodes to touch the beaker or each other. (See Fig. 1 below.) Note: Black wires are usually connected to the (-) terminal and red wires to the (+) terminal on the power supply. In electrolysis, the (-) terminal is called the cathode and the (+) is called the anode. The colors don't matter unless you are using it to keep track of polarity.
2. Obtain a power supply:
 - a) If you have a digital display, you do not need a multimeter to read the current (amps)—it will be displayed in the window. Connect one of the electrodes to the (+) terminal and the other to the (-) terminal of the power supply. Adjust the fine and course controls for current (amps) to their max setting. Adjust the voltage (volts) controls to the lowest setting.
 - b) If you have an analog display, you'll also need a multimeter to read the current (amps)—please follow the set up below. You'll need additional clamps to make it work—make sure they are not rusty.

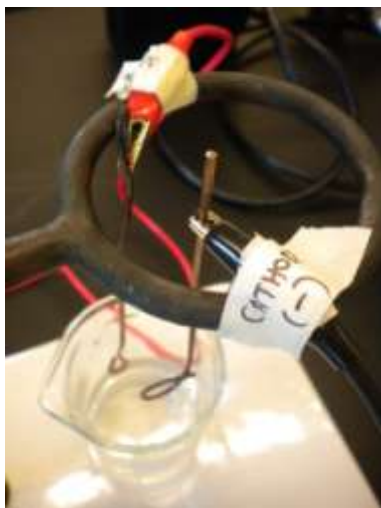


Figure 1. Electrodes and clips

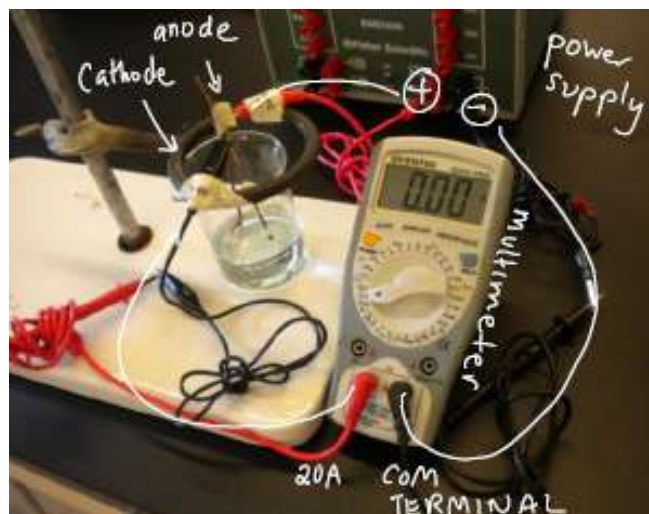


Figure 2. Electrolysis setup to include a multimeter.

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Set the proper voltage

(Skip this step if your power supply does not have an adjustable voltage)

3. Add about 50 mL of 0.50M sulfuric acid to your beaker and lower the ring clamp if necessary to immerse part of the electrodes in the acid. If you have a multimeter, set the dial to read DC (direct current), 20 or 20/20m for AMPS (not volts). If you have questions, ask your instructor.
4. Check your setup. Briefly turn the power supply on and increase the voltage until you see a reaction occurring inside the beaker. Adjust the voltage so that you have a very gentle stream of bubbles forming. The current should read somewhere between 0.1 and 0.5 amps (the lower the better). **Turn the power off as soon as you've set the voltage.**

If you do not see a reaction, check all of your connections and ask your instructor.

Start the experiment; Collect Data

5. Carefully remove the cathode and anode from your setup. Label one "anode" and the other "cathode" so you don't get them confused later. Rinse with acetone and when dry, weigh them to the nearest 0.001 g and record these masses—make sure you know which mass goes with each electrode. These are very important data!
6. Obtain a timer or watch to keep track of the length of your electrolysis.
7. Clamp the electrodes as before. Start the timer and the electrolysis at the same time. Record the current (Amps) every minute. If possible, adjust the voltage to keep the current fairly constant. There will be inevitable fluctuations in the current due to changes in the solution and temperature but your results will be better if the current is as constant as possible.
8. Collect data (Amps, every minute) for 5-10 minutes. Do two trials.
9. To stop the run, try to stop the stopwatch at the same time the power supply is switched off. Record the time elapsed, to the nearest second.
8. Carefully remove the anode and cathode. With extreme care, immerse it in 5ml of ethanol. **Do not wipe it.** Allow it to dry on a paper towel for a couple of minutes. Record the final mass of the anode and cathode.
9. Repeat once more, using new sulfuric acid and electrodes. (Clean your electrodes or use new electrodes if necessary).

Waste Disposal

When finished with both trials, put the sulfuric acid into a marked waste container.

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Collecting Data

Use this space to record data and observations. Remember, all data should be recorded in permanent ink.

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Calculations

For each, provide a calculation for both trials that you performed. Show your work for Trial 1 values.

REPORT AN APPROPRIATE NUMBER OF SIGNIFICANT FIGURES!!!

Q1) Calculate the average current during your experiment (in amps) and convert it to units of Coulombs per second. (1 Amp = 1 C/s)

Trial 1 = _____ C/s

Trial 2 = _____ C/s

Q2) Calculate the time that elapsed in seconds for each run.

Trial 1 = _____ s

Trial 2 = _____ s

Q3) Use the answers in Q1 and Q2 to find the total charge (in C) that passed through the circuit.

Trial 1 = _____ C

Trial 2 = _____ C

Q4) Use Q3 to calculate the number of electrons in each trial of the electrolysis. (One electron has a charge of 1.6022×10^{-19} C).

Trial 1 = _____ electrons

Trial 2 = _____ electrons

Q5) The copper atoms in the anode were converted to Cu^{2+} ions. Thus, **there were two electrons generated by each copper atom**. Based on the number of electrons generated in each trial (see Q4), calculate the number of copper atoms lost from the anode in each trial.

Trial 1 = _____ copper atoms

Trial 2 = _____ copper atoms

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Q6) Based on your mass measurements, how many grams of copper were lost from your anode* during each trial? (*If this data is compromised, you may use cathode data)

Trial 1 = _____ grams of Cu

Trial 2 = _____ grams of Cu

Q7) Based on Q5 and Q6, calculate the number of copper atoms lost from the anode per gram of copper lost from the anode

Trial 1 = _____ Cu atoms/gram

Trial 2 = _____ Cu atoms/gram

Q8) Use Q7 and the molar mass of copper to calculate the number of copper atoms per mole.

Trial 1 = _____ Cu atoms/mole

Trial 2 = _____ Cu atoms/mole

Q9) Take the average of the two trials in Q8. This is your measured value for Avogadro's number, N_A .

REPORT AN APPROPRIATE NUMBER OF SIGNIFICANT FIGURES!!!

N_A = _____ particles/mol

Q10) Calculate your percent error, which is the difference between your measured value and the accepted value (6.022×10^{23}) divided by the accepted value, multiplied by 100 and expressed as a positive value.

$$\% \text{ error} = \left| \frac{\text{experimental} - \text{accepted}}{\text{accepted}} \right| \times 100\%$$

% error = _____ %

Follow-Up Questions

1) Presumably, you used your anode data in your calculations. Try the calculations with your cathode data. Report your % error for Avogadro's number using cathode data.

2) How would the following sources of error affect your results? Would **your value** of Avogadro's number be **higher, lower**, or the **same** compared to the accepted value, 6.22×10^{23} ? Explain.

- a) You wiped your cathode at the end of the experiment, then weighed it post-electrolysis, using this mass to calculate Avogadro's number.

- b) During the course of the electrolysis, the current decreases significantly. You include all these data in your calculation of Avogadro's number.

- c) You used twice as much sulfuric acid to do your electrolysis.

- d) You did not allow your anode to dry completely before taking its mass at the end of the electrolysis.

Pre-Lab Assignment—To be completed BEFORE lab!

1. The thickness of a dollar bill is 0.15 mm. What is the height, in light-years of a stack of Avogadro's number of dollar bills? (Hint: 1 light-year is the distance that light travels in exactly 1 year. Light travels at a rate of approximately 2.998×10^8 m/s). Show your work and report your answers using significant figures.

2. (a) In your own words, what is a "mole" in chemistry?

(b) What are the units of Avogadro's constant, 6.022×10^{23} ?

3. (a) In the electrolysis, which electrode is attached to the (+) terminal? The (-) terminal?

(b) Why are the electrodes placed in sulfuric acid rather than pure water?

(c) How will you ensure that you don't confuse the anode and cathode during the experiment, especially when cleaning and weighing them?