AP Chem Lab: Electrolysis, Faraday and Avogadro Numbers

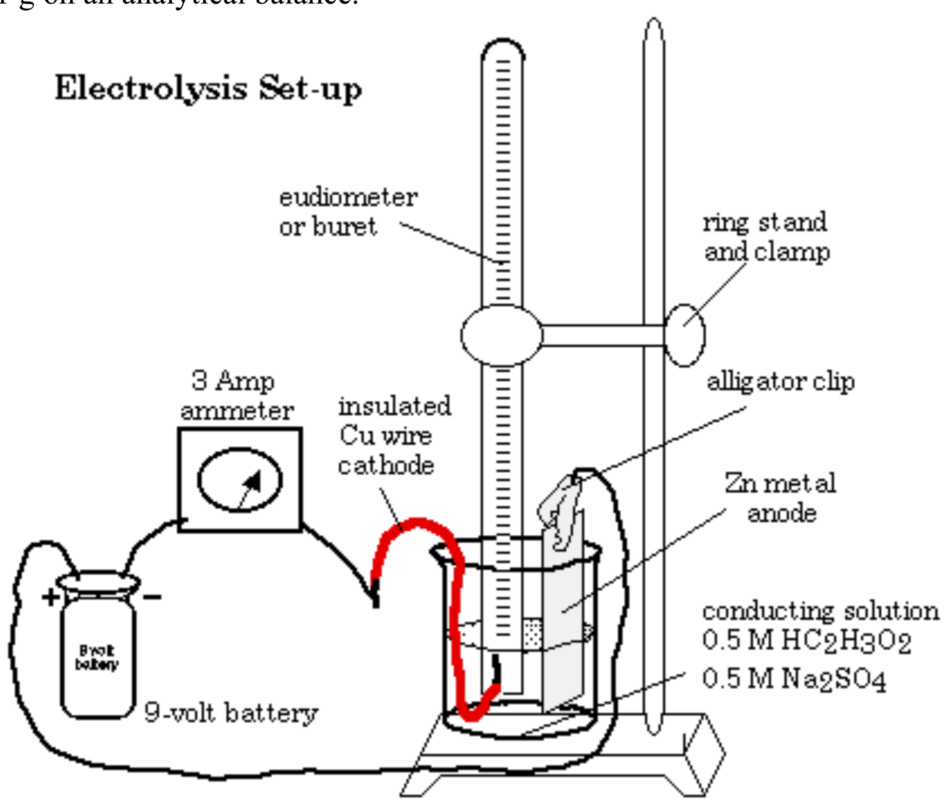
Background: Electrolysis is the process of using electricity to produce chemical change. In this activity, a device will be constructed in which electrolysis can be easily performed and observed. When an electric current is applied to an aqueous ionic solution, the positive and negative ions will move towards opposite electrodes. The positive ions will gain electrons from the negative electrode and try to undergo reduction. The negative ions will migrate to the positive wire and try to become oxidized by losing its excess electrons. Water may compete with the ions to be oxidized or reduced. If active metal electrodes are used for the experiment, they too may compete with the ions in solution to undergo oxidation. By comparing the standard reduction potentials, you can predict which species will be oxidized and which will be reduced.

A mole of electrons is a fundamental amount of electricity and is called a *faraday*. The accepted value for the amount of electrical charge for one faraday (1 mole of electrons) is 96500 coulombs. One coulomb is equal to the amount of charge that flows when 1 amp of current is transferred in 1 second (1 C = 1 amp•s). By quantitatively determining the amount of products formed by either the reduction reaction or the oxidation reaction or both, and measuring the average current in amps and the total elapsed time of the experiment, you can calculate the experimental value for a faraday *and* the value for Avogadro's number.

Problem: You are to separate a conducting solution by electrolysis, determine what products are formed during the oxidation and reduction steps, and calculate the experimental values for a faraday and Avogadro's number.

Procedure:

1. Obtain the materials needed to construct the electrolysis apparatus shown in the diagram below. If available, substitute a 2 amp/0-12 volt D.C. power source for the 9-volt battery. Lightly sand the zinc metal to clean it, rinse with water and then in acetone. Let the acetone evaporate to dryness, then mass on to ± 0.0001 g on an analytical balance.



2. Fill the eudiometer or buret completely with the conducting solution and invert into the beaker filled with the same solution. There should be about 100 mL of the conducting solution in the beaker.

3. The zinc metal will serve as the anode for the reaction. Connect the zinc to the positive pole of the battery with an alligator clip and immerse the metal (but not the clip) in the solution.

4. Insert the bare coiled end of the heavy copper wire up into the end of the buret. All but the bare end of the wire should be covered with watertight insulation. The copper electrode will be the cathode in the electrolysis cell. Connect the negative side of the battery to the negative side of the ammeter and then connect the positive side of the ammeter to the copper wire. Record the starting time for the experiment. Hydrogen gas bubbles will be generated above the copper wire and collect in the eudiometer.

5. Record the initial amp reading at the start of the electrolysis, and then every minute thereafter. Calculate the average flow of electricity in amps per minute.

6. Collect the gas until ~50 mL have been produced. At this point, stop the electrolysis by disconnecting the copper electrode from the power source. Record the elapsed time for the experiment.

7. Equalize the water levels inside and outside of the eudiometer, then record the exact volume of gas collected.

8. Measure and record the temperature of the solution and the barometric pressure during the lab. Look up the water vapor pressure at the experimental temperature in the CRC Handbook, or other reference source.

9. Discard the conducting solution down the drain. Rinse the zinc metal with 0.1 M acetic acid. Rub off any loose adhering coating with a paper towel. Then rinse the metal with water and then acetone. Let the acetone evaporate, and mass the metal to ±0.0001 g.

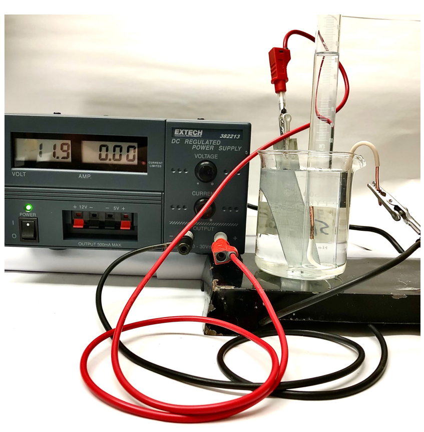
10. Calculate the total coulombs of electricity transferred during the elapsed time of the experiment.

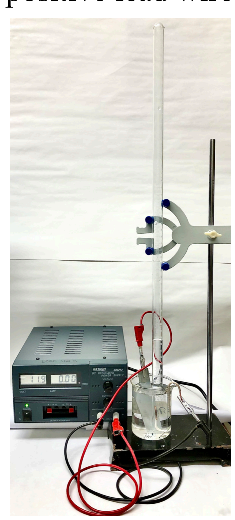
11. Use the Ideal Gas Law to calculate the moles of hydrogen gas produced. Use this quantity to calculate the value for 1 faraday and the value for Avogadro's number.

12. As a comparison, calculate the moles of zinc metal that was oxidized, then use this value to also determine the value for 1 faraday and for Avogadro's number. Determine the percent error for both methods and discuss sources of error and the accuracy of each method.

Results & Conclusions: Consult a reduction potential table to determine the half-reactions and write net equations for the changes that occur during the electrolysis reaction. Show all calculations for determining the moles of hydrogen gas formed and the moles of zinc oxidized, as well as the values for Faraday's constant and Avogadro's number. Include an error analysis of the accuracy of the experiment and explain sources of error.

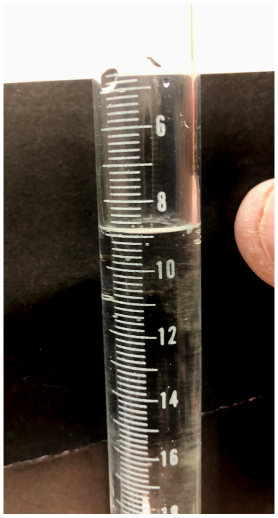
Modifications: As an alternative power source, a lab power supply with a rating of 2 amps/0-12 D.C. volts, can be used. Adjust the knob of the power supply to a voltage setting where a reaction begins to occur. Two 9-volt batteries can be connected in parallel by twisting together both positive lead wires and then both negative lead wires to approximately double the amps produced.



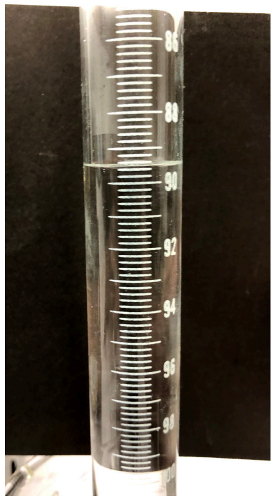


Sample Experimental Data: Electrolysis, the Faraday and Avogadro Numbers

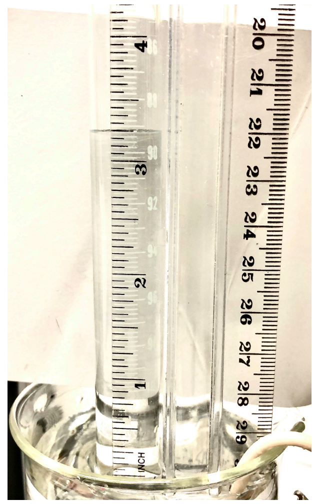
Intial volume in eudiometer



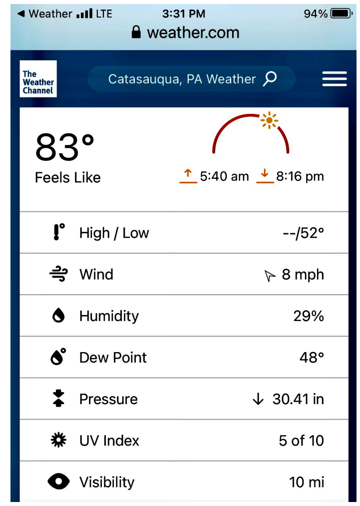
Final volume in eudiometer



Difference in water levels



Weather Channel experimental conditions



By using the power supply, the volume of hydrogen gas was collected in the eudiometer over a time period of 13.0 minutes with an average current reading of 0.80 amps.

Water Vapor Pressure Based on Experimental Temperature (32°C)

Vapor Pressure of Water from 0 °C to 100 °C

T °C	P (torr)	T °C	P (torr)	T °C	P (torr)	T °C	P (torr)	T °C	P (torr)
0	4.6	17	14.5	35	42.2	53	107.2	71	243.9
1	4.9	18	15.5	36	44.6	54	112.5	72	254.6
2	5.3	19	16.5	37	47.1	55	118.0	73	265.7
3	5.7	20	17.5	38	49.7	56	123.8	74	277.2
4	6.1	21	18.7	39	52.4	57	129.8	75	289.1
5	6.5	22	19.8	40	55.3	58	136.1	76	301.4
6	7.0	23	21.1	41	58.3	59	142.6	77	314.1
7	7.5	24	22.4	42	61.5	60	149.4	78	327.3
8	8.1	25	23.8	43	64.8	61	156.4	79	341.0
9	8.6	26	25.2	44	68.3	62	163.8	80	355.1
10	9.2	27	26.7	45	71.9	63	171.4	81	369.7
11	9.8	28	28.4	46	75.7	64	179.3	82	384.9
12	10.5	29	30.0	47	79.6	65	187.5	83	400.6
13	11.2	30	31.8	48	83.7	66	196.1	84	416.8
14	12.0	31	33.7	49	88.0	67	205.0	85	433.6
15	12.8	32	35.7	50	92.5	68	214.2	86	450.9
16	13.6	33	37.7	51	97.2	69	223.7	87	468.7
		34	39.9	52	102.1	70	233.7	88	487.1

Intial mass of zinc metal used

Final mass of zinc metal after electrolysis

Check out the video that demonstrates the set up of the lab equipment and sample data generated during the experiment at https://youtu.be/gHZa0lC_xqM