

Determination of Faraday's Constant

Minneapolis Community and Tech. College
C1152 Principles of Chemistry II
v.5.10

Introduction

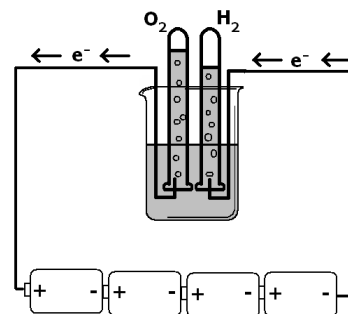
Recent increases in the price of non-renewable energy (oil, natural gas, and gasoline) and global warming concerns have motivated a search for energy alternatives. One of these is hydrogen gas produced via the electrolysis of water. This reaction, shown below, requires energy to decompose water molecules into hydrogen and oxygen gas.



This is to say that hydrogen is not a principle fuel source but rather a means of storing energy for later use. The energy required to create hydrogen and oxygen gas must come from some other electrical source like wind, solar or nuclear energy. Although using solar energy to produce hydrogen gas is very environmentally advantageous (sunlight → electricity → hydrogen), it is an expensive alternative in comparison to other energy sources due to the inefficiencies and costs of solar cells.

Once produced, the hydrogen gas must be stored in either high pressure reinforced gas tanks or as a condensed liquid in a thermally insulated container (fancy thermos bottle). In either case, the storage requirements require specialized, expensive equipment very different from what is currently used to store liquid fuels like ethanol and gasoline. Consequently, before hydrogen can be used as a fuel globally it will be necessary to address its specialized storage requirements.

In the lab, the electrolysis of water is a relatively simple experiment requiring only acidified water, inert platinum electrodes, gas collection containers and a source of electricity. The diagram at right shows the apparatus. Note that two moles of H_2 are produced for each mole of O_2 . Consequently, the volume of H_2 collected will be twice the O_2 volume. Also note that the platinum electrodes are un-reactive and therefore will not be oxidized or reduced.



Determination of Faraday's Constant

In today's experiment, you will determine the value of Faraday's constant. Faraday's constant is the amount of charge carried by 1 mole of electrons. Since 1 electron has a charge of 1.602×10^{-19} coulombs (C), one mole of electrons should have a charge of....

$$1 \text{ F} = \left(6.02214 \times 10^{23} \right) \times \left(1.60218 \times 10^{-19} \text{ C} \right) = \mathbf{96485 \text{ C}} \quad (\text{i.e. } 96485 \text{ C/mole})$$

To experimentally determine a value for Faraday's constant, we'll need to know both the total amount of charge (Q) in Coulombs and the number of moles of charge (n_e). We can then calculate a value for Faraday's constant

$$F = Q/n_e \quad \text{(Equation 1)}$$

The total charge (Q) is calculated by knowing the electrical current (I) and the total time (t) for the current flow. Electrical current (I) is the amount of charge that flows through a wire per second (i.e. Coulombs per second). Mathematically:

$$I_{(\text{amperes})} = Q_{(\text{Coulombs})} / t_{(\text{seconds})} \quad \text{(Equation 2)}$$

Therefore, we can calculate the total charge by multiplying current by time:

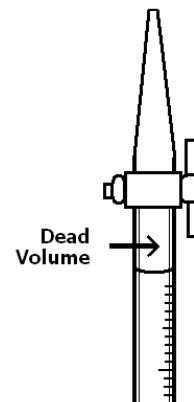
$$Q_{(\text{Coulombs})} = t_{(\text{seconds})} \times I_{(\text{amperes})} \quad \text{(Equation 3)}$$

In this experiment, you will keep the current constant at 250 mA for 12 minutes.

The total number of moles of electrons (n_e) are obtained by carefully collecting the hydrogen gas that is produced as a result of the electrolysis of water, specifically, the reduction of 2H^+ into H_2 (Cathode Reaction):



The hydrogen gas will be collected using the water displacement technique. As the H_2 is produced, it will bubble up into an inverted burette. The volume of the gas is determined by reading the gas level on the side of the burette and adding to it the burette's dead volume (the space that isn't on the burette's graduated scale).



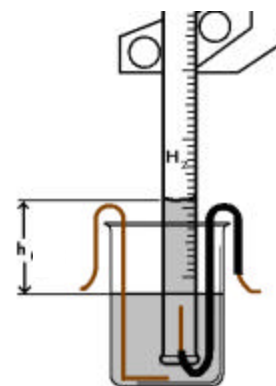
The pressure of the gas inside the burette would be ideally atmospheric pressure P_{atm} . However, the presence of water vapor reduces H_2 's contribution to the total pressure as does the weight of the water suspended in the tube. Thus the actual pressure of the hydrogen gas is calculated as follows:

$$P_{\text{H}_2} = P_{\text{atm}} - P_{\text{H}_2\text{O Vapor}} - P_{\text{suspended}} \quad \text{(Equation 4)}$$

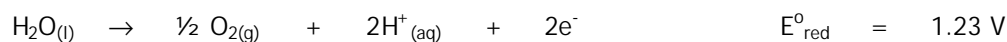
where P_{atm} is the atmospheric pressure obtained from the barometer, $P_{\text{H}_2\text{O Vapor}}$ is the vapor pressure of water (Table 5.3 pg. 197 of your textbook...4th ed) and $P_{\text{suspended}}$ is found by first measuring the height of the suspended H_2SO_4 solution (see figure at right) in millimeters and using the following conversion:

$$P_{\text{suspended}} (\text{torr}) = \text{height (mm)} \times 0.07355923 (\text{torr/mm}) \quad \text{(Equation 5)}$$

Use the ideal gas law, the total gas volume and adjusted pressure you determined above to calculate the number of moles of gas formed (n_{gas}). Determine the number of moles of electrons, n_e using n_{gas} and the cathode reaction above.



Anode Reaction: Because inert platinum electrodes are NOT used in this experiment, *two oxidation reactions are possible at the anode:*



However, only ONE can occur. *You should decide which oxidation reaction will occur based on its position in the standard reduction potential table before coming to class.*

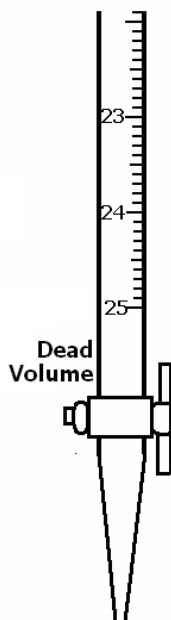
Procedure

Required Materials

- 1 ... Rubber squeeze bulb
- 1 ... 25 mL buret
- 1 ... buret clamp
- 2 ... 250 mL beaker (cell & H₂SO₄ storage)
- 1 ... ring stand
- 1 ... 10.00 mL volumetric pipette
- 1 ... Pipette helper
- 1 ... Low voltage power supply
- 1 ... Vernier current sensor
- 1 ... Stopwatch
- 1 ... Barometer
- 1 ... ruler (mm)
- 1 ... ~ 12 inch wire banana-banana
- 2 ... ~ 12 inch wire banana-alligator clip
- ~175 mL ... 3 M H₂SO_{4(aq)}

Burette Dead Space Determination

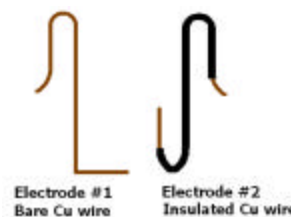
We will be using an inverted burette to collect the H₂ gas produced in this experiment. However, before we can use the burette to measure the volume of the gas produced we must first accurately determine the amount of "dead space" in the region between the burette's valve and the last scale marking (see figure at right). Do this by first pipetting 10.00 mL of distilled water (10.00 volumetric pipette) into the burette (make sure the valve is closed). Read the burette and subtract this number from 25.00 mL. This result is the amount of the original 10.00 mL that can be obtained by reading the scale of the burette. Subtract this number from 10.00 mL to obtain the "dead volume".



Later you will use the burette's scale to obtain the volume of the H₂ gas produced. Add the "dead volume" to this volume to determine the total volume of the gas.

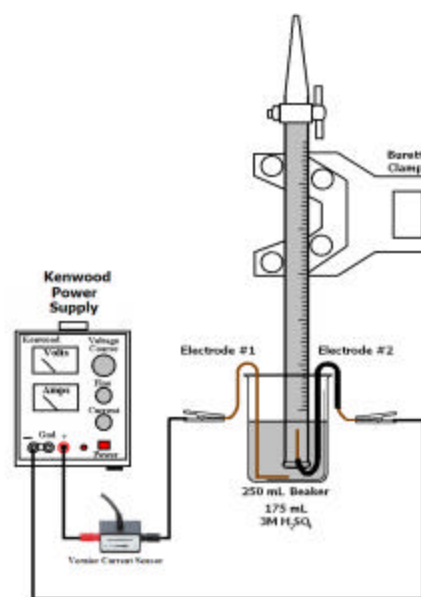
Electrode Construction:

Construct the copper wire electrodes shown at right. Electrode #1 is made from bare copper 12 gauge wire bent such that it can hang over the edge of a 250 mL beaker. Another bend is made



so that an approximate 3 cm section of wire sits flat on the bottom of the beaker. Electrode #2 is constructed from *insulated* 12 gauge wire. Remove approximately 2 cm of the insulation from each end of the wire. This electrode should also be bent so that it hangs on the side of the beaker opposite Electrode #1. Note that the part of Electrode #2 inside the beaker is *bent vertically upwards*. The bend should be made at a point where the wire is insulating (look carefully at the diagram to see this). The vertical segment will be positioned within the inverted burette during the experiment.

Electrolytic Cell



Position the two electrodes in the beaker as shown in the figure above. **STIR THE MIXTURE WITH A STIR BAR AND STIR PLATE.** Obtain a clean 25 mL burette, burette clamp and ring stand. Position the burette, up side down, in the burette clamp and lower it into the beaker such that the exposed metal of Electrode #2 is completely contained within the open end of the burette. Pour approximately 120 mL of 3 M H₂SO₄ into the beaker (*use the scale on the side of the beaker*).

Double check the positions of the two electrodes to be **SURE** they are **NOT TOUCHING** each other.

Make absolutely sure the valve of the burette is turned off before beginning the next step.

Squeeze the air from the suction bulb and attach it to the top, pointed end of the buret. Carefully/slowly open the burette valve just far enough to permit H_2SO_4 to be drawn up into the burette. Turn the valve off when solution completely fills the burette from top to bottom. *Don't draw solution up into the vacuum bulb.*

Wiring

You will be provided with three pieces of wire. One of the wires has "banana" plugs on each end. Use this to connect the **positive terminal** of the power supply to the **red terminal** of the Vernier Current sensor.

The other two wires are equipped with banana plugs on one end and alligator clips on the other. Use one wire to connect the Vernier Current Sensor's **black terminal** to Electrode #1 using the alligator clip. Use the remaining wire to connect the power supply's **negative terminal** to Electrode #2.

Power Supply and Multi-meter Settings

Initially set the power supply knobs as follows:

- Voltage Coarse Control:
FULL CCW (counter clockwise)
- Voltage Fine Control: Midway
- Current: FULL CW (clockwise)

Plug in the Vernier current probe to port #1 of the data logger. Plug in the power supply and turn it on. Slowly turn up the voltage on the power supply with the Voltage Coarse knob. Stop when the computer display reads approximately 0.250 A. Turn off the power supply using the power switch (leave the Voltage Coarse knob set).

During this time, H_2 bubbles should have been forming at Electrode #2 and collecting in the burette. If you see gas forming at Electrode #1 the wires connecting the electrodes are backwards. Reverse the wires before proceeding further.

Refill the burette with H_2SO_4 using the suction bulb as you did earlier. (i.e. remove the H_2)

Electrolysis Experiment

Begin stirring the mixture at 600 speed.

Obtain a stopwatch and zero it. Be sure that the burette is completely full of H_2SO_4 . When you and your partner are ready, simultaneously start the stop watch and turn on the power supply. Use the power supply's "FINE" adjustment to keep the current sensor's reading as close to 250 mA as possible throughout the experiment. If dramatic changes are necessary, return the "FINE" control to the midway position and adjust the "Coarse" control.

Your goal is to keep the electrical current constant (0.250 A) during an entire run.

Check to be sure hydrogen gas is being collected in the burette and not bubbling into the atmosphere. If you determine it necessary to make adjustments, refill the burette and start over.

Turn the power supply off when exactly 12 minutes have passed. Determine the volume of gas collected by making the necessary burette scale readings. Don't forget to include the "dead volume" volume you determined earlier.

Measure the distance between the liquid level in the burette and the liquid level of the beaker. This measurement will be used to adjust the gas pressure for the downward pull or weight of the solution in the burette.

When you are ready, have the instructor collect the hydrogen gas you've generated.

Repeat the experiment a second time.

Cleanup: Carefully pour the used sulfuric acid into the waste container provided. Rinse all apparatus with distilled water. Be careful not to splash acid or rinse solutions on you or your clothing.

Acknowledgements

This experiment is based on an experiment designed by Dr. Victor Singleton, El Paso Community College, El Paso, Texas.

V. Team Report

Page 1:

- **Upper right corner:** Your name, Your lab section number, Date of experiment
- **Data Table (below)**

		Trial 1	Trial 2
Lab Temperature	°C		
Barometric Pressure	torr		
Electrical Current	Amperes		
Time	seconds		
Q Total Charge (Equation 3)	Coulombs		
Burette Dead Space	mL		
Volume _{H₂} Measured	mL		
Volume _{H₂} Adjusted	mL		
P _{vapor H₂O} (Table 5.3)	torr		
h _{solution}	mm		
P _{suspended} (Equation 5)	torr		
P _{H₂} (Equation 4)	torr		
n_{gas} (ideal gas law)	moles		
n_e moles of e ⁻ transferred **	moles		
F (Equation 1) <i>Faraday's Constant</i>	C/mole		
Δ% Compare F to the known value.			

Page 2:

Answers to the following questions:

1. What is the half cell reaction that occurs at Electrode #1. What observations support your answer?
2. What changes would be observed if Electrode #1 was made from platinum wire instead of copper?
3. Why was it necessary for electrode #2 to be insulated?