

## Determination of Faraday's Constant

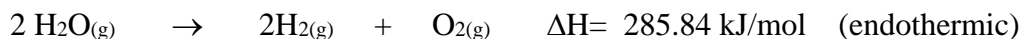
Minneapolis Community and Tech. College

C1152 Principles of Chemistry II

v.2.16

### I. Introduction

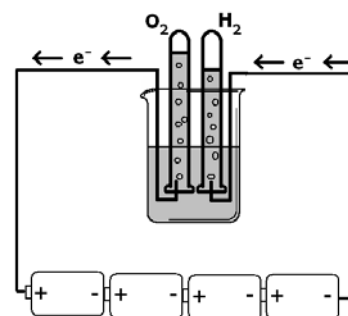
Recent energy issues and global warming concerns have motivated a search for energy alternatives. One of these is hydrogen gas produced via the electrolysis of water. This reaction, shown below, requires energy to decompose water molecules into hydrogen and oxygen gas.



Hydrogen is not a principle fuel source like fossil fuels, but instead a means of storing energy for later use. If the energy required to create hydrogen and oxygen gas comes from some other electrical source like wind or solar energy, it could be considered a "clean" process. Although using solar energy to produce hydrogen gas is very environmentally advantageous (sunlight  $\rightarrow$  electricity  $\rightarrow$  hydrogen), it is currently an expensive and inefficient alternative in comparison to other energy sources.

Once produced, the hydrogen gas must be stored in a compressed form using high pressure gas tanks or as a condensed liquid in a thermally insulated containers (a fancy thermos bottles). 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 for general consumption, 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  $\text{H}_2$  are produced for each mole of  $\text{O}_2$ . Consequently, the volume of  $\text{H}_2$  collected will be twice the  $\text{O}_2$  volume. Also note that the platinum electrodes are un-reactive and therefore will not be oxidized or reduced.



### Determination of Faraday's Constant

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

$$1 \text{ F} = \overset{\text{1 mole of } e^-}{(6.02214 \times 10^{23})} \times \overset{\text{charge of 1 } e^-}{(1.60218 \times 10^{-19} \text{ C})} = \mathbf{96485 \text{ C}}$$

Thus, Faraday's constant is known as 96485 C/mol.

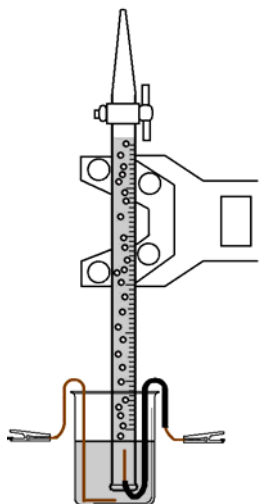
To experimentally determine and confirm the value for Faraday's constant, the total amount of charge (**Q**) in Coulombs and the corresponding number of moles of charge (**n<sub>e</sub>**) are required. Faraday's constant is then determined by:

$$F_{\text{exp}} = \mathbf{Q/n_e} \quad (\text{Equation 1})$$

## Determination of $n_e$

The moles of electrons ( $n_e$ ) value is determined from the number of moles of hydrogen gas,  $n_{H_2}$ , collected during an experimental trial. Calculating  $n_{H_2}$  requires we use the ideal gas law and the known temperature (T), volume (V) and pressure (P) of the  $H_2$  gas:

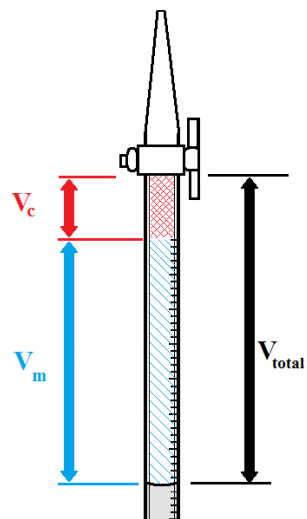
$$P_{H_2} V_{total} = n_{H_2} RT \quad \text{(Equation 2)}$$



**$H_2$  Volume ( $V_{total}$ ):** Hydrogen gas is collected via water displacement (figure at left). As  $H_2$  is produced at the cathode, gas bubbles are collected inside an inverted burette initially filled with sulfuric acid solution. As the  $H_2$  gas is collected, the level of the liquid drops inside the burette.

The  $H_2$  gas collects in in two different parts of the burette (figure at right). The first region ( $V_c$ : **Constant volume**) is the space between the valve and where the burette markings begin. Even though there are no markings for this region, its volume can be experimentally determined and is the same for each subsequent trial.

$H_2$  gas also collects in a second region ( $V_m$ : **Measured Volume**) where there are burette markings to facilitate this measurement. Care must be taken to determine the volume correctly as the measurement scale is upside down. It is good to double check the  $V_m$  measurement by actually *counting the number of mL of gas in this region*.



The actual volume of the gas inside the burette ( $V_{total}$ ) is determined by adding together  $V_c$  the and  $V_m$  contributions:

$$V_{total} = V_c + V_m \quad \text{(Equation 3)}$$

**$H_2$  Pressure ( $P_{H_2}$ ):** The pressure of the hydrogen gas inside the burette would be ideally be the pressure of the atmosphere,  $P_{atm}$ . 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 below the gas. Thus, the actual pressure of the hydrogen gas is calculated as:

$$P_{H_2} = P_{atm} - P_{H_2O} - P_{sus} \quad \text{(Equation 4)}$$

T, °C	P, torr
0	4.5851
5	6.5450
10	9.2115
15	12.7931
20	17.5424
25	23.7695
30	31.8439
35	42.2037
40	55.3651
45	71.9294
50	92.5876

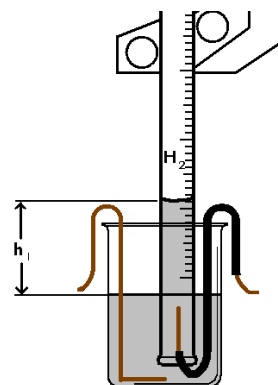
[https://en.wikipedia.org/wiki/Vapor\\_pressure\\_of\\_water](https://en.wikipedia.org/wiki/Vapor_pressure_of_water)

The vapor pressure of water,  $P_{H_2O}$  is determined from the known vapor pressures and temperatures provided in the table at left. For temperatures between those listed on the table, it is necessary to interpolate. Linear interpolation is sufficient for these purposes.

$P_{sus}$  is a corrective term that adjusts for the weight of the suspended liquid pulling down on the gas column. This effect would be non-existent if the liquid level in the burette matched the liquid level in the beaker. As this is often not the case, we calculate the correction  $P_{suspended}$  in torr as:

$$P_{sus} \text{ (torr)} = h \text{ (mm)} \times 0.073555923 \quad \text{(Equation 5)}$$

This correction requires the height of the suspended solution column,  $h$ , in millimeters (Figure at right) which is measured with a plastic ruler.



Once the ideal gas law is used with values for P, V and T, to determine  $n_{H_2}$ , moles of electrons ( $n_e$ ) is calculated utilizing this reduction reaction that takes place at the cathode:



## Determination of Q

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 in units of Coulombs/second. Mathematically, we can determine the rate of flow of charge by dividing the charge by time:

$$I_{(\text{amperes})} = Q_{(\text{Coulombs})} / t_{(\text{seconds})} \quad (\text{Equation 6})$$

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

$$Q_{(\text{Coulombs})} = t_{(\text{seconds})} \times I_{(\text{amperes})} \quad (\text{Equation 7})$$

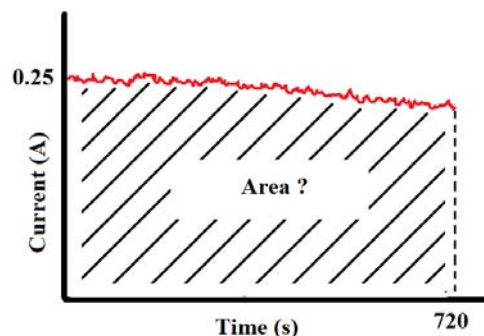
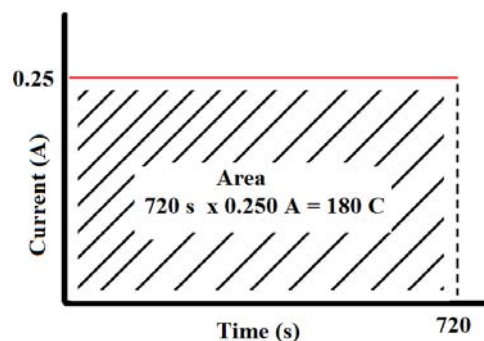
In this experiment, you will adjust the power supply for  $I = 0.250 \text{ A}$  and the experiment will go on for 12 minutes (720 seconds). Ideally, if we were to graph the electrical current vs. time, we'd have the graph at right. Note that as long as the electrical current is constant, the result of our  $Q_{(\text{Coulombs})} = t_{(\text{seconds})} \times I_{(\text{amperes})}$  calculation is simply the area of the rectangle;

$$\begin{array}{rclcl} 720 \text{ s} & \times & 0.250 \text{ A} & = & 180 \text{ Coulombs} \\ \text{Length} & \times & \text{Height} & = & \text{Area} \end{array}$$

In reality, the electrical current fluctuates due to increases/decreases in  $\text{H}^+$  concentration near the cathode will also increase or decrease. Whilst vigorous stirring helps even out these effects, it cannot entirely eliminate the current variations.

The figure at right illustrates typical current fluctuations observed in this experiment. Note also that the current decreases slightly throughout the trial. Nonetheless, the area under the curve represents the total charge that flows over the course of the experiment. However, determining the area isn't as straightforward as multiplying the length and width of a rectangle.

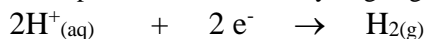
In calculus courses you'll learn that the area under a curve can be determined by calculating the "integral" of the function. In our case, no calculus is needed as the Vernier software has the ability to calculate the area beneath a curve using what's called numerical integration. To accomplish this, the computer divides up the area under the curve into very small squares and then counts them up.



To determine the area under the curve (Q), first select all the data by dragging the mouse cursor over the graph. Then click on the "integral" button (figure at left). The result of the area calculation (integral) is displayed on the graph in a text box and that value is recorded as the total charge Q for the experimental trial.

## II. Prelab Exercise... Clearly answer these questions in INK in your lab notebook before coming to lab.

- How many moles of electrons are required to produce 15 moles of hydrogen gas via the reduction reaction?

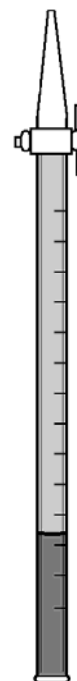


- Determine water's vapor pressure at  $28.5^\circ\text{C}$  using the data table above.

- The burette at right was used to collect  $\text{H}_2$  gas (lightly shaded). If  $V_c = 2.1 \text{ mL}$ , what is the total volume of the gas? Note that this burette has 1 mL increments. The burettes used in the actual experiment has 0.1 mL increments.

- Whilst performing this experiment, the friendly chemist accidentally bumps the power supply and the current increases from 0.25 A to 0.32 A for several seconds. She then readjusts the current to the previous 0.25 A level and continues the experiment. Why is it *not necessary* to re-start the experiment from the beginning?

- Where does the energy come from for this experiment?



### III. Word Processed Report


Page 1: Upper right hand corner Name, Lab section number and date

#### Data Table

Page 2: Answers to the following questions:

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

IV. Data table (Don't round until  $F_{\text{exp}}$ )  $R = 0.082057 \text{ L atm mol}^{-1}\text{K}^{-1}$



		Trial 1	Trial 2
<b>T</b>	°C		
<b>P<sub>atm</sub></b>	torr		
<b>Q</b>	Coulombs		
<b>V<sub>c</sub></b>	mL		
<b>V<sub>m</sub></b>	mL		
<b>V<sub>total</sub></b>	mL		
<b>P<sub>H<sub>2</sub>O</sub></b>	torr		
<b>h</b>	mm		
<b>P<sub>sus</sub></b>	torr		
<b>P<sub>H<sub>2</sub></sub></b>	torr		
<b>n<sub>H<sub>2</sub></sub></b>	moles		
<b>n<sub>e</sub></b>	moles		
<b>F<sub>exp</sub></b> <i>Round Appropriately Here</i>	C/mole		
<b>Δ% Compare F<sub>exp</sub> to 96485 C/mol</b>			

## V. Procedure

### V<sub>c</sub> Determination

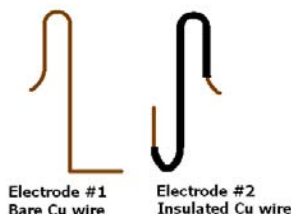
Pipette 10.00 mL of distilled water using a 10.00 volumetric pipette into the burette (make sure the valve is closed).

Read the burette and subtract this number from 25.00 mL. This 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 V<sub>c</sub>. V<sub>c</sub> is typically in the 1 – 3 mL range.

### Copper Electrodes:

The electrodes used in this experiment are shown at right.

Electrode #1 is the anode where oxidation occurs. It is made from bare copper 12 gauge wire bent such that it can hang over the edge of a 150 mL beaker.

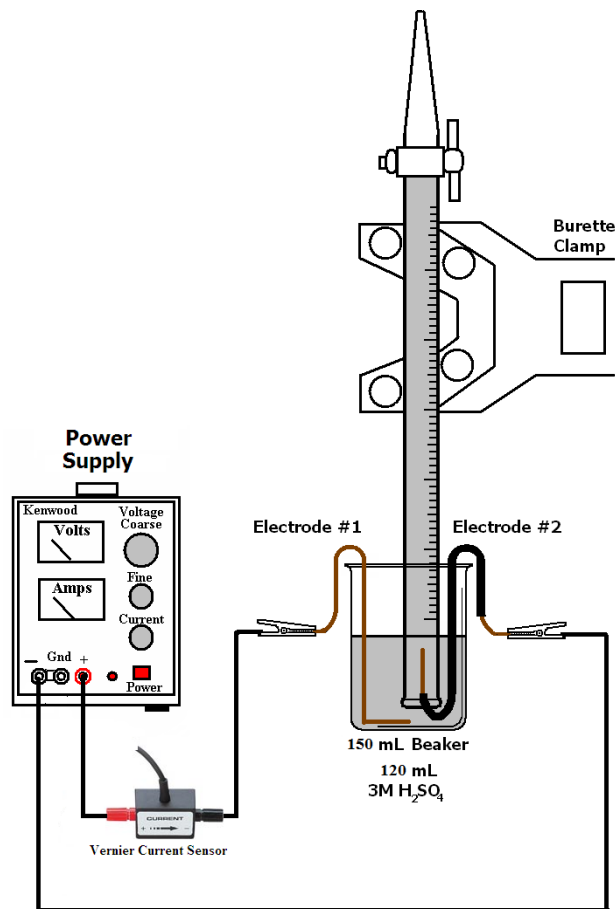


Another bend is made so that a circular loop of wire sits flat on the bottom of the beaker.

Electrode #2 is the cathode where reduction occurs. It is constructed from *insulated* 12 gauge wire and also 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 vertical segment should be positioned within the inverted burette during the experiment.

## Experimental Setup



Position the two electrodes in the beaker as shown in the figure above.

Locate the 25 mL burette, upside down in the burette clamp and lower it into the beaker so that the exposed metal of Electrode #2 is completely contained within the open end of the burette.

Place the magnetic stirring bar in the beaker and make sure it can spin freely without bumping into either electrode. For best results, position the beaker in the exact center of the stir plate.

Pour approximately 120 mL of 3 M H<sub>2</sub>SO<sub>4</sub> into the beaker (*use the scale on the side of the beaker*).

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 burette. Carefully/slowly open the burette valve just far enough to permit H<sub>2</sub>SO<sub>4</sub> 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 wires. 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 (don’t change any knob settings).

H<sub>2</sub> bubbles should have been forming at Electrode #2 and collecting in the burette during this test run. 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<sub>2</sub>SO<sub>4</sub> using the suction bulb as you did earlier. (i.e. remove the H<sub>2</sub>)



## Electrolysis Experiment

Stir the mixture at 600 rpm.

Be sure that the burette is completely filled with H<sub>2</sub>SO<sub>4</sub>.

When ready, simultaneously click on the “collect button” and turn on the power supply using its power switch.

Check to be sure hydrogen gas is being collected in the burette and not bubbling into the atmosphere. You should also see the electrical current vs. time graph begin to fill in.

Data collection will automatically stop when 12 minutes have passed. **Make sure to turn off the power supply at that time using the power button.**

Measure the distance “h” between the liquid level in the burette and the liquid level of the beaker using a plastic ruler and record this in your notebook.

“Q” determination. Select all of your graphical data by dragging your mouse cursor over it. Now, click on the integral button. The area under the electrical current curve will be displayed on the graph and is equal to “Q”. Record this value in your notebook.

When you are finished, 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. (You may want to consider wearing an apron)

## Acknowledgements

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