

Faraday's laws proven correct through a lab experiment of electrolysis

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During the course of my experiment, I shall attempt to prove Faradays two laws regarding electrolysis correct. I will attempt to keep my experiment accurate within the limits of human error. Faradays Laws

1The mass of a substance produced at an electrode during electrolysis is proportional to the quantity of electricity that has passed.

The number of Faradays required to discharge one mole of an ion at an electrode is equal to the number of charges on that ion. My Investigation In order to investigate the Faradays laws I must measure the amount of metal collecting on the cathode, or more accurately the amount of metal lost by the anode in the experiment that I intend to perform, whilst varying the current passing through the solution and keeping the time constant thereby changing the number of coulombs passing through the solution as: $\text{Coulombs} = \text{Current} * \text{Time}$

I will be able to show whether or not the amount of metal deposited at the cathode (or the amount lost by the anode) is proportional to the quantity of electricity that has passed by plotting a graph of one value against the other, i. e. metal produced against number of Coulombs. This will prove or disprove the first of Faradays laws within the limits of experimental error.

As: 1 Faraday = 96500 C/mol And this is the required amount to discharge 1 mole of electrons or charged ions, by varying the current and thereby varying the number of Coulombs. By plotting a graph of number of coulombs against the amount of metal removed from the anode, I will be able to see how many coulombs are required to discharge a certain amount of metal in grams. By converting the number of grams into the number of moles and <https://assignbuster.com/faradays-laws-proven-correct-through-a-lab-experiment-of-electrolysis/>

dividing the number of coulombs by the number of moles, and dividing the value by two (as copper has a charge of 2+) I will be able to calculate a value for the Faraday and therefore by comparing it to the actual value for the Faraday, test whether or not this value is correct within the limits of experimental error. Experiment Plan I will use a solution of copper sulphate, and although the quantity should make little difference to my results, I will use the same volume (300 cm³) of solution.

I will take an electrolysis non-rimmed beaker and using clamp stands to prevent the electrodes touching, I will place two pure copper electrodes in the solution and run a current through the same solution to cause copper to be lost from the anode and deposited on the cathode. I will run the current through the solution for a time of ten minutes for each experiment. I will record both the current passed through and the mass before and after electrolysis of both the anode and cathode. I will weigh both the anode and cathode before and after the experiment and record the results, I will also clean the electrodes using emery paper so that the surfaces are clean and the copper sticks and is lost easily. I will also clean the electrodes using a process of washing them in water, ethanol and propanone in order to remove any loose pieces of copper from the surfaces of the electrodes and making them as clean as possible. Prediction

I predict that in my experiment, I will find that when I calculate the number of coulombs by multiplying the current passed through by the time taken, and then plot this against the amount of metal lost at the anode, I will find a good positive correlation and a direct proportionality in accordance with

Faradays 1st law. I also predict that when I plot the number of coulombs against the amount of metal removed from the anode, a similar direct proportionality will be shown. I also predict that when I divide the number of coulombs by the number of moles of metal lost, I will find that it comes to roughly one Faraday or 96500 C/mol. My Equipment I shall set up my equipment as shown below: Fair Test In order to keep my experiment as fair and accurate as possible:

- I will keep the copper electrodes the same throughout all my experiments
- I will use the same copper sulphate solution throughout all my experiments
- I will use the same volume of copper sulphate throughout all my experiments
- I will keep the time period constant at 10 minutes throughout all my experiments.

I will clean the electrodes carefully using emery paper, water, ethanol and propanone before and after each experiment and before weighing the electrodes.

| Results | Anode | Cathode | Anode | Cathode | Anode | Cathode | Anode | Cathode |
|---------------------|-------|---------|-------|---------|-------|---------|-------|---------|
| Initial Mass | 93.23 | 94.86 | 92.35 | 95.44 | 95.49 | 93.35 | 92.07 | 95.90 |
| Final Mass | 92.99 | 95.04 | 92.07 | 95.90 | 95.23 | 93.53 | 91.63 | 96.41 |
| Current (A) | 1.00 | 1.00 | 1.50 | 1.50 | 2.00 | 2.00 | 2.50 | 2.50 |
| Mass gained or lost | 0.24 | 0.18 | 0.28 | 0.46 | 0.26 | 0.18 | 0.44 | 0.51 |
| Analysis | 0.55 | 0.71 | | | | | | |

In order to analyse my results I have drawn a graph of current against mass lost at the anode, as I feel that the mass lost at the anode is more accurate than the mass gained at the cathode. This is due to the fact that some of the copper may be lost in the transition between the electrodes, through it not sticking to the cathode and therefore making the results inaccurate. Although my results are not entirely accurate, I have been able to find a general trend giving me a straight line

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through the origin. This shows me that the current is directly proportional to the amount of metal removed from the anode, similarly as I kept the time period constant at 10 minutes I can show that the number of Coulombs is directly proportional to the amount of metal removed from the anode by electrolysis.

By plotting a graph of number of Coulombs against the amount of metal removed from the anode, I will be able to calculate the direct proportionality between the number of Coulombs passed through the solution and the amount of metal removed from the anode. I will take some readings from my graph in order to calculate the direct proportionality: When number of coulombs = 800 Then number of grams of copper removed from anode = 0.245 When number of coulombs = 1000 Then number of grams of copper removed from anode = 0.305 When number of coulombs = 1200 Then number of grams of copper removed from anode = 0.367 When number of coulombs = 1400 Then number of grams of copper removed from anode = 0.429 When number of coulombs = 1600 Then number of grams of copper removed from anode = 0.490 When number of coulombs = 1800 Then number of grams of copper removed from anode = 0.550 In order now to investigate Faradays second law, using number of grams in one mole of copper. $\text{Cu} = 64$, therefore, number of moles removed from the anode during electrolysis = $x/60$ $0.245/60 = 0.00408$ and $800/0.00408 = 196078$ $0.305/60 = 0.00508$ and $1000/0.00508 = 196850$ $0.367/60 = 0.00612$ and $1200/0.00612 = 196078$ $0.429/60 = 0.00715$ and $1400/0.00715 = 195804$ $0.490/60 = 0.00817$ and $1600/0.00817 = 195838$ $0.550/60 = 0.00916$ and $1800/0.00916 = 196507$

The correct value for the above is two Faradays, which is equal to 193000. My average value of 196193 seems to be a little too high. I will attempt to explain this problem in my evaluation. However, it does show me that the error is constant using my values from the graph and therefore, there is a constant number, the Faraday, that relates the number of Coulombs required to remove one mole of metal removed from the anode. Conclusion

My Experiment has proved that the amount of metal removed from the anode is directly proportional to the current passing through the solution by virtue of a straight line through the origin when time, t , was kept constant. My results did not exactly fit the straight line, however, the correlation between the points was very close to the straight line and the error shown in the results that were anomalous can be attributed to experimental error. This proves Faradays First Law.

The second law was more difficult. After plotting a careful graph of the number of Coulombs against the number of grams of copper, I found that several of my results were anomalous. However, I was able to plot a reasonable line of best fit due to the obvious correlation between the points and hence take readings from the graph. However, on attempting to calculate the Faraday, or number of Coulombs per mole by converting grams of metal into moles of metal and then by dividing the number of Coulombs by the number of moles of metal, I attempted to find the number of Coulombs per mole, this should be equal to 193500 or 2 Faradays. However, I found that my results showed a constant error and were consistently too large, this is due to my own experimental error, and I will expect to explain

this constant error in my evaluation. However, this experiment has shown that the Faraday is equal to 193500 as my results are close enough to this value to show Faradays second law. Evaluation

Although on the whole my trends are good and my results prove Faradays laws to be correct within the limits of experimental error, my experiment was far from perfect and many of my results were inconsistent. I have also found a uniform error in the calculation of the value of one Faraday. The inconsistency in the results and the uniform error can be explained in the following manner: The weighing of the electrodes may have been inaccurate. It makes a difference how the electrodes are placed on the balance as it can alter the reading. The discrepancy in the readings could be accounted for in this way. In the weighing of the electrodes, there was the potential for a ± 0.01 g error in each reading that could cause a minor discrepancy in the results although this is relatively insignificant. The current readings on the electronic ammeter had a tendency to vary during the experiment depending upon the resistance of the solution; this led to a ± 0.05 Amp error in the results that were obtained. This alteration in the current gives rise a potentially large difference between the current passing through the Copper Sulphate solution and the current recorded.

The cleaning of the electrodes may not have been as good as I thought it was. Although I attempted to clean the electrodes as much as I could, it is possible that I did not clean them well enough and therefore prevented the copper sticking to the cathode as well as it should have done. This would explain the copper residue observed under the cathode during the

experiment. The cleaning process may have loosened some copper that fell from the anode during electrolysis making my results inaccurate. This would explain the small amounts of copper observed underneath the anode as a residue during the electrolysis. As my results were not all the same and several points did not fit the general trend, I was forced to draw a line of best fit for analysis purposes. This again brings a large amount of potential for error into the analysis, as a line of best fit can only ever be an estimate. I have attempted to explain the reasons for my anomalous results above. The line of best fit could also account for the consistent error found when investigating Faradays 2nd Law.

However, despite my results needing some adjusting before analysis, my results successfully proved Faradays laws within the limits of experimental error. However, if I were to repeat the experiment, I would spend more time cleaning and weighing the electrodes, and attempt to keep the current more constant by use of a purer solution of copper sulphate. I also could extend my investigation by finding the relative atomic mass of copper and calculating the percentage error in this value, thereby working out my experimental error.