The Charge of an Electron

An Indirect Measurement of the Charge of an Electron in terms of Avagadro's Number

Purpose: This very brief experiment that will allow you to measure the charge e of an electron.

Apparatus: A beaker, 0.2 molar $CuSO_4$ solution, copper plates (2), rubber bands, and a DC power supply.

Discussion: The idea of this experiment is very simple. We use an electrolytic cell and a DC power supply to deposit copper onto one of the plates in the electrolytic cell. As the electro plating process ensues we record the time and the current that is flowing. (A constant current supply will ideally be provided by your lab instructor. If none is available your instructor will advise you how to proceed) An ampere is one coloumb flowing past a point in the wire per second, and so the product of the time (in seconds) we have had the current flowing in the cell and the current (in amps) is the number of coloumbs of charge that have coursed through the wires and therefore through the electrolytic cell.

$$Q = It \tag{3.1}$$

Then, by weighing the amount of of copper that has been deposited we will know how many moles of electrons have been through the cell. (A mole of copper weighs 63.5 grams) Thus, using Avagadro's number (*i.e.* 6.02 x 10^{23}) we will be able to compute the charge (in Coloumbs) per electron.

Before outlining the procedure let us describe the operation of the electrolytic cell. In essence, an electrolytic cell consists of two conducting plates placed in a container holding a solution (See Figure1). The solution in this experiment is simply water in which some copper sulfate has been dissolved. When a "salt" such as $CuSO_4$ is dissolved in water it dissasociates into ions, in this case, Cu^{+2} and SO_4^{-2} ions. As "solvated ions" (that is, ions covered with water molecules) they are essentially non-interacting and move rather freely throughout the liquid.

When a D.C. voltage V is put across the cell there arises an electric field $\vec{E} = V/d$ between the electrodes (in this case copper plates 1 and 2). (See Figure 2) The ions respond to the force $\vec{F} = q\vec{E}$ and drift in this applied electric field towards the plate of the sign opposite their own. This motion of charges inside the liquid constitutes a current equal in magnitude to that which is flowing in the circuit. This is due to the fact that when the ions, say the Cu^{+2} ions get to the negative plate they steal two electrons from that plate and then lay as reduced (metallic) copper, Cu, on or near that electrode. Two electrons flow in from the rest of the circuit to replace those two that were absorbed by the Cu ion.

A similar process occurs at the other electrode. There SO_4^{-2} ions collect and donate two electrons to the plate. In the process, chemical reactions occur at the plate which cause gas (SO_2) to be evolved from that electrode and cause and several other processes at that plate occur which cause the plate to erode, that is, causes some of the copper on that plate to go into solution.

The basic point is that for each Cu atom that is reduced on the plate, two electrons will have flowed through your circuit.

Safety Precautions

1) NEVER TOUCH THE ELECTRODES WITH WET HANDS WHILE THE D.C. VOLTAGE IS ON. Just shut off the cell and proceed. The entire experiment happens at pretty low voltage BUT even at low voltage, if you are a good enough conductor, you may indeed get a nasty shock.

2) DON"T SNIFF THE (rather small amount of) GAS THAT IS EVOLVED BY THE PLATES AND DO THAT PART IN THE HOOD. Or atleast please be sure that you do this in a reasonably

well ventilated room. Open the window. It's a small amount of gas and so it shouldn't be any problem. BUT don't consciously try to inhale it!

3) TAKE CARE IN HANDLING THE BEAKER AND THE SOLUTION. Ordinary common sense precaution is in order here.

Procedure: First prepare the plating solution by making a solution with these ratios: 125g of $CuSO_4$*$5H_2O$ to 900mL DI and (while stirring) 35mL H_2SO_4 and 50mL of Ethyl Achohol. You may want to make enough for a second try as well.

Then clean and dry the copper electrodes that are provided. After scrubbing them a bit with iron wool and wiping them with a cloth it is best to place them in an over/on a hot plate at about 150-180^{o}C or so for about ten minutes to drive off any residual moisture that may be sticking to them, and to let it oxidize just a little bit. Carefully pick them up in a paper towel (don't touch them!) and go up to the very precise balance provided in the lab (by the sink) and record their weight.

Now, back in the optics lab, assemble the electrolytic cell. This is most easily done by forming the sheet into cylinder cathode (-) THen make spiral anode (+) of clean copper wire wound tight so it fits inside the cylindrical cathode. Attach wires to each and insert the anode inside the cylinder and suberge it in the solution. Try to do so without bending/scraping the plates and without getting them too marked up with your fingerprints. When you are done binding them be sure that the plates are not touching one another, that is, not making electrical contact and that furthermore the spacing between the cylinder and the wire spiral is somewhat uniform. This completes the construction of the cell electrodes.

IMPORTANT: The electrodes need not be totally submerged. We strongly recommend that the contact points of the plates with the wires are well above the surface of the solution.

Now decide what the current level should be: an appropriate current density should be in the 0.02A/cm^2 range. Record current you are shooting for and the time (about 20 minutes?) to get good change in the mass (the deposited copper). Now we are ready to energize the cell. WITH THE D.C. SUPPLY OFF hook up the cell and the DVM as shown in Figure 1. Note that now the DVM should be set to AMPS. Now, recording the start time, you need to turn on the D.C. Watch the ammeter intently OR record the current on a digital device.

Now disconnect the apparatus, retrieve the cylinder and try to avoid dropping any of the plated copper, the reddish sand that you will see clinging to the cathode. When done dip if you can dip the cathode in DI and Alchohol and Ether if you have some. Place it in/on the oven/hot plate again for a few minutes or so to let them dry. Record the new weight of that electrode, the positive one with the deposited Cu.

After weighing the electrodes, etc dry, throw away the Cu dust, clean and wash the electrodes and return to them to the lab. and clean up a bit,m making sure that the chemicals are properly stored. DO NOT RETURN THE $CuSO_4$ SOLUTION THAT YOU USED TO THE 0.2 MOLAR , but just to the "used" jar of $CuSO_4$ solution.

Conclusions:

Now that we have made all our measurements, we are ready to compute our measured value for the electron's charge. Since we recorded the current and the time that it flowed in the cell we have (by Eq.(3.1)) a measure for the total number of coloumbs of electricity that have been transported through the solution. Now, comparing the weight of the (-) electrode before and after we know how many moles of copper were deposited on the plate. Thus we can compute the number of copper

atoms that represents (using Avagadro's number) and from there compute the charge of a Cu^{+2} ion (in coloumbs) by dividing the total charge by the number of copper atoms.

Q: What is then your measurement for the charge of an electron?

Try to estimate the errors in your measurements of the current and the time as well as the weighings. Combine them to get a quantitative estimate of the error in your measurement of the electron's charge.

Q: What systematic effects can you think of that effect the measurement?