Current as a Flow of Charge (Electroplating)

Purpose

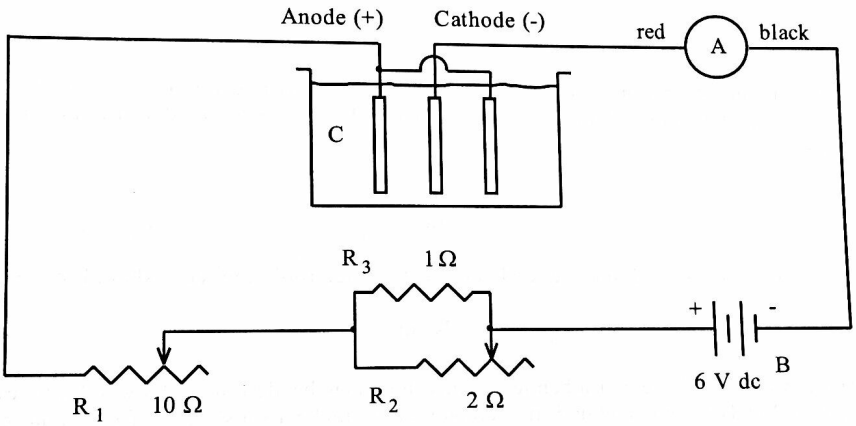
The purpose of this experiment is to study current as a flow of charge by observing the action of an electroplating cell. It is also desired to determine the value of Faraday's constant and the electrochemical equivalent of copper.

Theory

As an example of the action that takes place when a current is sent through a chemical cell, consider the circuit below. A power supply (B) is in series with an ammeter (A), a chemical cell (C), a rheostat (R_1), and a parallel combination of a rheostat and fixed resistor (R_2 and R_3). The chemical cell has two outer copper plates that serve as the anode and a center copper plate that serves as the cathode. These copper plates are immersed in a solution of copper sulfate ($CuSO_4$).

Figure 1

Circuit Diagram for Use with the Electroplating Cell



Some of the copper sulfate is dissociated into doubly-charged positive copper ions Cu^{++} and doubly-charged negative sulfate ions SO_4^{--}. Under the action of the electric field between the electrodes, the copper ions move toward the cathode (−), acquire two electrons, and deposit themselves on the cathode as copper atoms. At the copper anode (+), copper atoms are relieved of electrons and go into solution as doubly-charged positive ions. The net effect of the action in this circuit is that copper atoms are taken from the anode and deposited on the cathode, with the copper ions moving through the solution while the electrons move through the external metallic conductors. As much copper is deposited on the cathode as is removed from the anode; and the solution remains unchanged.

If M is the total mass of copper deposited on one electrode in time t by some current I, then

$$M = KIt \tag{1}$$

where K is a constant of proportionality known as the electrochemical equivalent of the substance. The constant K is also identified by the relationship

$$K = \frac{A}{FV} \tag{2}$$

where A represents the atomic weight of the element deposited, F is Faraday's constant, and V is the valence of the ion.

Given a particular type of electroplating cell, the current I in amperes, and the time t in seconds, it is possible to predict the mass M of the metal plated onto the cathode as follows:

$$Q = I t \tag{3}$$

Equation (3) gives the amount of charge in coulombs transferred around the circuit and follows from the definition of current as the amount of charge transferred per unit time. The charge transferred through the electroplating cell in a given period of time must be equal to that transferred through the ammeter and to that transferred through any other part of the circuit. Otherwise, charge would rapidly accumulate at some point in the circuit, would exert a repulsive force on approaching moving charges, and would very quickly stop the flow of current. In other words the current flowing in any part of a series circuit is the same.

The number of electron charges N_e transferred through the electroplating cell is given by equation (4), where e is the absolute magnitude of the charge of one electron.

$$N_e = \frac{Q}{e} \tag{4}$$

The number of atoms N_a transferred through the electroplating cell is given by equation (5). V is the number of electron charges per ion and is called the valence.

$$N_a = \frac{N_e}{V} \tag{5}$$

The mass m of one atom can be determined from equation (6), given the atomic weight A and Avogadro's number N_o. One mole of an element has a mass of A grams by the definition of the mole; and it contains N_o atoms.

$$m = \frac{A}{N_o} \tag{6}$$

Finally, the total mass deposited on the cathode of the electroplating cell should be as given by equation (7).

$$M_{pred} = N_a m \tag{7}$$

The measured value of the electrochemical equivalent can be determined by solving equation (1) for K. Since Q has already been calculated, it is convenient to make a substitution from equation (3).

$$K_{meas} = \frac{M_{meas}}{Q} \tag{8}$$

The accepted value of K can be determined from a handbook of chemistry or physics or by substituting the accepted values of A, F, and V into equation (2). The measured value of Faraday's constant can be determined by solving equation (2) for F and substituting the measured value of the electrochemical equivalent.

$$F_{meas} = \frac{A}{K_{meas} V} \tag{9}$$

It is interesting to note that Faraday's constant can be evaluated in terms of more-fundamental constants. First substitute the value of m from equation (6) in equation (7).

$$M = \frac{N_a A}{N_o} \tag{10}$$

Then substitute the value of N_a from equation (5) into equation (10).

$$M = \frac{N_e A}{N_o V} \tag{11}$$

Then substitute the value of N_e from equation (4) into equation (11)

$$M = \frac{Q A}{N_o e V} \tag{12}$$

Then substitute the value of Q from equation (3) into equation (12).

$$M = \frac{A}{N_o e V} It \tag{13}$$

But this is the same form as the equation obtained by combining equations (l) and (2).

$$M = \frac{A}{F V} It \tag{14}$$

By comparing constants in equations (13) and (14), it is seen that

$$F = N_o e \tag{15a}$$

or

$$N_o = \frac{F}{e} \tag{15b}$$

Faraday's constant can be measured with an electroplating cell, and the charge of the electron can be measured by means of Millikan's oil drop experiment, enabling Avogadro's number to be calculated from equation (15b).

Apparatus

Copper-plating cell
Copper sulfate solution
Rheostat, 10 ohm with 2 amp fuse
Rheostat, 2 ohm
Resistor, 1 ohm
Digital multimeter
Analytical balance
Isopropyl alcohol
Emery paper
Stopwatch

Procedure

1. Make the following check on the digital voltmeter before connecting it into the circuit: The AC-DC button should be in the out position. Depress the "Ma" button and the "2000" button. Plug the power cord into a 120 Vac outlet and depress the "POWER" button. The display should read 0000 (or at worst 0001). If it does not, call the laboratory instructor. Plug the black lead into the "COMMON" jack of the multimeter and the red lead into the "Ma" jack.

2. Connect the apparatus according to the schematic drawing shown in Figure 1. Use the center plate of the copper-plating cell as the cathode and the outer pair of plates as the anode. Have the laboratory assistant check your wiring before proceeding.

3. Fill the copper-plating cell about three-quarters full of copper sulfate solution.

4. Turn the knob on the 10 ohm rheostat fully counter-clockwise. Set the knob on the 2 ohm rheostat to its center position. Plug the power supply into a 120 Vac outlet, turn its power switch on, and adjust the knob on the power supply for a voltage of roughly 6 volts. Adjust the 10 ohm rheostat for a current of approximately 1 ampere. (The power supply knob and the other rheostat may also be used if necessary.)

5. Turn off the power supply and remove the center plate from the copper plating cell. Clean the plate with fine emery paper and rinse it with water and alcohol. Be sure your hands are clean and dry when handling the plate; and hold it only by the edges.

6. When the plate is dry, weigh it to the nearest tenth of a milligram on the analytical balance. Follow the instructions of the laboratory instructor carefully in using the analytical balance. It can be easily damaged. Record the initial mass of the plate on the data sheet.

7. Reinsert the plate into the cell without touching the portion of the plate to be immersed. Clamp it firmly in place.

8. Turn on the power supply and note the exact time of closing. Use the rheostats to keep the current constant for approximately 30 minutes. Note the exact time at which the circuit is broken. Record the value of the current in amperes and the time interval in seconds.

9. Lift off the top of the cell and remove the center plate carefully. Rinse it first in clean water, then in alcohol, and air dry without rubbing.

10. Weigh the plate to the nearest tenth of a milligram, using the same analytical balance as for the initial weighing. Record the final mass on the data sheet. Subtract the initial mass to find the measured mass of copper deposited on the center plate (M_{meas}).

11. Calculate Q, N_e, N_a, M_{pred}, K_{meas}, and F_{meas} by using equations (3) through (9). You may use an Excel spreadsheet to do these calculations quickly in the lab. But be sure to show all these calculations in the Calculations section of your laboratory report.

12. Compare M_{pred} with the measured increase in mass of the center plate. Find the percent difference. Compare F_{meas} with the accepted value of Faraday's constant. Find the percent error.

Questions

1. How can an electroplating circuit function as a current measuring device?

2. Why are two anodes used instead of one?

Experiment EM-2 Data Sheet

Experimental measurements:

Mass of center plate originally:	25.8282	g
Mass of center plate finally:	26.437	g
Mass of copper deposited (M_{meas}):	0.6088	g
Electric current (I):	1	A
Time (t):	1800	s

Standard constants required:

Atomic mass of copper (A):	63.546	g/mole
Valence of copper (V):	2	
Absolute magnitude of charge of electron (e):	1.602×10^{-19}	C
Avogadro's number (N_o):	6.022×10^{23}	parts/mole
Mass of one copper atom (m):	1.055×10^{22}	g
Faraday's number (F)	96.485	C/mole

Calculated results:

Charge transferred (Q):	1800	C
Number of electron charges transferred (N_e):	1.124×10^{22}	
Number of atoms transferred (N_a):	5.618×10^{21}	
Predicted mass of copper deposited (M_{pred}):	0.5928	g
Percent difference between M_{pred} and M_{meas}:	2.659	%
Measured value of electrochemical equivalent of copper (K_{meas}):	3.382×10^{-4}	g/C
Measured value of Faraday's number (F_{meas}):	93.941	C/mole
Percent error in measurement of Faraday's number:	2.636	%