Experiment 21E

FV 3/29/05

ELECTROCHEMISTRY: CELLS, BATTERIES, CORROSION, AND ELECTROLYSIS

MATERIALS: Ag, Cu, Zn strips; sandpaper; Mg ribbon; Cu wire; 0.25 M solutions of $AgNO_3$, $Cu(NO_3)_2$, $Zn(NO_3)_2$, 1.0 M HCl; $CuSO_4 \cdot 5H_2O$ (s); concentrated NH_3; plastic disposable pipets; 0.5 M NaCl; 400 mL beaker with 3.6 M H_2SO_4; Pb strips; J-electrodes with Pt foil; double-throw switch; DC power supply; various electrical leads; digital multimeter; 1.0 M H_2SO_4; 0.20 M KI, graphite rods; 0.0200 M $Na_2S_2O_3$; phenolphthalein in dropper bottles; graduated gas collection tube; starch solution; light bulb apparatus; 50-mL buret; thermometer; paper strips for salt bridges; 1.5 M KNO_3; MnO_2; saturated NH_4Cl solution; carbon block; square wooden block; zinc plate; filter paper; 100 mL, 250 mL, 400 mL and 600 mL beakers.

PURPOSE: The purpose of this experiment is to illustrate the principles and practical aspects of galvanic cells, commonly used batteries, corrosion, and electrolysis.

LEARNING OBJECTIVES: By the end of this experiment, the student should be able to demonstrate the following proficiencies:

1. Determine the half-reactions involved in spontaneous oxidation-reduction reactions.
2. Build galvanic cells and measure their electrical potentials and currents.
3. Draw schematic representations of cells, labeling electron flow direction, anode, cathode, and ion migration directions from salt bridges.
4. Use standard reduction potential tables and the Nernst Equation to predict the expected electrical potentials for cells.
5. Explain how atmospheric corrosion occurs, and how it results in the eating away of metal.
6. Explain how concentration cells produce an electrical potential.
7. Explain the function of the lead storage battery and the dry cell battery.
8. Determine the likely anode and cathode reactions in an electrolysis experiment.
9. Apply Faraday's Law to an electrolytic cell, verifying the stoichiometric relationships implied by the half-reactions.

DISCUSSION:

General chemistry texts provide significant coverage for many aspects of electrochemistry, including the activity series, oxidation-reduction reactions, galvanic cells, batteries, electrolysis, and corrosion. Some specific issues that are often not addressed as thoroughly in texts are discussed below:

Simple wiring diagrams. Electrochemistry involves simple electrical circuits, instruments for measuring electrical potentials and currents, electrodes, and external power supplies for pumping electrons through in certain directions, etc. Diagrams are used to depict these kinds of circuits and devices. Sophisticated diagramming conventions have been developed and will undoubtedly be used in later courses in physics and electrical engineering. The diagrams shown below are simple and are intended to mainly illustrate the connectivity of the various components of the circuits used in this experiment.

Simple galvanic cell circuit. Figure 1 depicts the simple circuit that allows for the determination of the electrical potential and/or current generated by a simple galvanic cell. The details of the electrodes are not included in this diagram.

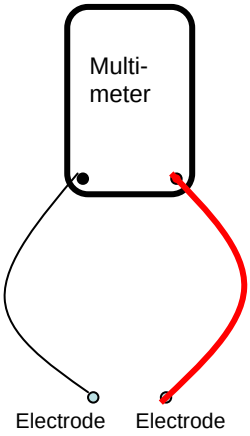

Figure 1. The wiring diagram for a simple galvanic cell connected through a multimeter. The electrodes may be such things as strips of metal immersed in solutions. The circuit is only completed if there is a way for electric current to flow between the electrodes, which is the reason that salt bridges are often required in such cells. As discussed below, the multimeter is capable of measuring either the electrical potential or the electric current generated by the cell. For the multimeters used in this experiment, when the electric potential reading is positive, the red wire will be attached to the cathode.

Circuits involving the double-throw switch. Figure 2 depicts two circuits which can be selected by throwing a special type of switch known as the double-throw switch. In the particular case shown, one of the circuits is a simple galvanic cell circuit, similar to that shown in Figure 1, and the other circuit is an electrolytic cell circuit, in which an external power supply pumps electrons in a certain direction.

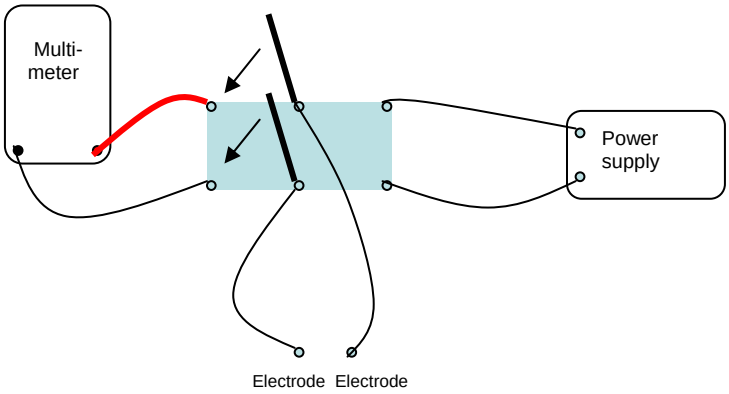

Figure 2. In this diagram, the double throw switch is about to be closed towards the left. Once closed, the circuit will be identical to that in Figure 1. If the switch were thrown to the right, the circuit would include the power supply, as in an electrolytic cell. A multimeter can also be included in this right-hand circuit if one desires to monitor the electric current during electrolysis.

How multimeters work. Multimeters are electronic instruments which can measure several different electrical properties, e.g., electrical potential, current, and resistance. We will primarily be concerned with electrical potential and current in electrochemical applications.

When the multimeter is set to one of the voltage scales, it acts as a potentiometer. A potentiometer measures the voltage difference between two points of a circuit by creating an equal (but opposite) electromotive force and applying it to the circuit until the current is reduced to zero. Thus, when a multimeter is set to one of the voltage scales, there is no current flow, and the electrochemical process being investigated is actually stopped, even though the electric potential generated by the cell is measured. The standard unit for electrical potential is the volt, V.

When a multimeter is set to one of the current (amps) scales, the electrochemical process is not stopped by any opposing forces, as in the case when electrical potential is being measured. Instead, the current produced by the process is easily measured by applying appropriate physics principles,

e.g., the relationship between electric current and magnetic fields. The standard unit for electric current is the ampere, which is defined as 1 coulomb of electric charge per second.

Passivation. Many metals, especially active metals, will react with components of the environment, e.g., oxygen gas in the air, to form a surface layer of metal oxide. In some cases, the layer is very tightly bound to the underlying pure metal, and may prevent further attack of the metal. This phenomenon is known as "passivation." Aluminum, magnesium, and zinc all form passivating layers of oxides which provide varying degrees of protection to the underlying metal. Iron oxides, on the other hand, are not tightly bound to the underlying metal, and easily flake away and expose the underlying iron to further attack. Passivating layers of metal oxides are often very thin and can usually be removed by immersing the electrode in strong acid. Due to the amphoterism exhibited by aluminum, its passivating layer will also be removed by immersion in strong base. In many cases, the passivating layers can be removed mechanically through sanding. Some depassivated metals, such as aluminum, reform the oxide layer very quickly due to the high reactivity of the metal with oxygen.

Real cells have real issues. General chemistry texts treat galvanic cells at an introductory level, without going into some of the details that are unavoidably associated with such cells. For example, varying degrees of efficiency in the motions of ions as they pass from the salt bridge into other cell compartments can reduce the effectiveness of the electromotive force, hence yielding measured voltages somewhat lower than expected. In addition, for those metals which tend to form passivating layers of metal oxides, the measured cell potentials are somewhat complicated, even if the bare metal has been exposed through sanding or some other chemical means, since the passivating layers tend to reform rather quickly. Even though the electrons can pass through ultra-thin passivating layers, the electrical potential generated in this situation is different than when the passivating layer is completely absent. The consequence of these kinds of complicating issues is that unless great care is taken to address them, the measured cell potentials will differ somewhat (generally, they will be lower) from those predicted by using the tabulated half-reaction potentials and the Nernst Equation.

Another issue that unavoidably affects measured electrical potentials is the ionic strength of the solutions in electrochemical cells. As has been noted in other experiments several times, ions in solution have an impact on the behavior of each other in such a way that the simple relationships often used in general chemistry are not precisely valid. This is especially true when ion concentrations are above 0.010 M. Some of the deviations of measured electrical potentials for cells can be attributed to these effects, especially when the concentrations of ions are in the 1 M range or higher, as is the case with many of the cells and batteries in this experiment.

Atmospheric corrosion. Atmospheric corrosion occurs when dissolved oxygen is reduced at the cathode of an electrochemical cell. In the absence of other reducible chemical species, dissolved oxygen is a viable candidate for reduction, via the half-reaction

$$O_2(aq) + 2\ H_2O + 4e^- \rightarrow 4\ OH^- \qquad (1)$$

If a metal sufficiently high on the activity series, such as iron, is electrically coupled to a region in contact with dissolved oxygen, the metal will undergo oxidation (i.e., it will corrode). For the case of iron, the oxidation reaction is

$$Fe(s) \rightarrow Fe^{2+} + 2e^- \qquad (2)$$

Even if the anode and cathode regions of a cell are made from the same type of metal, atmospheric corrosion will occur if there is a nonuniformity if the dissolved oxygen concentration, such as occurs with the hull of a ship where the dissolved oxygen concentration is greater near the water surface.

These concentration differences are enough to cause the different regions of the same piece of metal to act anodically or cathodically. For more information about naval applications of electrochemistry and corrosion, go to the website http://www.chemistry.usna.edu/navapps/PDF/Corrosion.pdf .

Electrolysis. General chemistry textbooks include detailed information about electrolysis, including applications of Faraday's Law, and should be referred to for more background information.

Lead Storage battery. General chemistry textbooks provide detailed information about lead storage (also known as lead acid) batteries, and should be consulted in preparation for this experiment.

Most textbook discussions about lead storage batteries omit information about the process known as the "initial charging" of the battery. This is the process by which the lead storage battery is prepared for discharge, followed by recharging, processes which are discussed in detail in most texts. The "initial charging" process is simply the electrolysis involving two lead strips immersed in sulfuric acid. The strips are connected to an external power supply which pumps electrons from one strip to the other, forcing oxidation reactions occur at the anode, and reduction reactions to occur at the cathode. Textbooks provide guidance about determining which of the possible redox reactions will likely occur at the anode and cathode during electrolysis. For the case of the lead storage battery undergoing "initial charging," the candidate materials for oxidation include Pb(s) and H_2O, while candidate materials for reduction include H_2O and H^+(aq). Standard reduction potential tables include reactions involving these materials, from which it is possible to select the likely anode and cathode reactions during electrolysis. For the "initial charging" of a lead storage battery, the possible anode reactions are:

$$2\ H_2O \rightarrow O_2 + 4\ H^+ + 4\ e^- \qquad E^o = -1.23\ V \qquad (3)$$
$$Pb(s) + 2\ H_2O \rightarrow PbO_2 + 4\ H^+ + 4\ e^- \qquad E^o = -0.665\ V \qquad (4)$$

The possible cathode reactions are:

$$2\ H_2O + 2\ e^- \rightarrow H_2(g) + 2\ OH^- \qquad E^o = -.83\ V \qquad (5)$$
$$2\ H^+ + 2\ e^- \rightarrow H_2(g) \qquad E^o = 0.00\ V \qquad (6)$$

Dry Cell battery. Unlike the lead storage battery, which is rechargeable, the dry cell is known as a primary cell (i.e., it cannot be recharged). This cell has a zinc cylinder as the anode and a carbon rod as the cathode. The carbon rod is surrounded by a paste formed by mixing solid MnO_2 with an aqueous solution of NH_4Cl. The zinc cylinder is kept separate from the paste by a porous paper, called the separator, so that there is no direct contact of the reactants associated with the anode and cathode. The electrode reactions are:

Anode $$Zn(s) \rightarrow Zn^{2+}(aq) + 2e^- \qquad (7)$$

Cathode $$2\ MnO_2(s) + 2\ NH_4^+(aq) + 2\ H_2O + 2e^- \rightarrow 2Mn(OH)_3(s) + 2\ NH_3(aq) \qquad (8)$$

The function of the salt bridge is effectively performed as the Zn^{2+} ions migrate away from the anode, passing through the paper separator, and complex with the NH_3 produced at the cathode, forming $Zn(NH_3)_4^{2+}$. In this way, charge balance is maintained in both the anode and cathode regions of the cell, and neither product builds up concentration near the electrode where they are produced, which would, according to the Nernst Equation, quickly diminish the cell's electrical potential. The porous separator only allows the dissolved aqueous ions to pass through, keeping the Zn(s) and the MnO_2(s) from coming into direct contact.

Electrical energy and power. When an electric circuit operates, charged particles pass through regions with different electrical potentials. The energy change for a particle of charge q, passing through an electric potential difference E, is qE. (1 J = 1 Coulomb x 1 Volt) Thus, the total amount of electrical energy produced by an electrochemical cell is equal to the total charge passing through the circuit of the cell, multiplied by the potential difference generated by the cell. Energy-producing devices are also characterized by a power rating, i.e., the amount of energy produced by the cell per unit time. Since current is the amount of charge passing a point per unit time, the power output of a cell is simply the current multiplied by the potential difference generated by the cell. (1 W = 1 J/sec = 1 Amp x 1 Volt)

PROCEDURE:

1. Your instructor will demonstrate the setup of a galvanic cell, including the simple paper salt bridge, the use of the double-throw switch, and the use of the digital multimeter. They will also discuss the need for removing oxide layers from the underlying metals (e.g., via sanding) before use in electrochemical applications. Make any notes for future reference, if needed.

Activity Series and Simple Galvanic Cells

2. Add 5.0 mL each of 0.25 M $Cu(NO_3)_2$, 0.25 M $Zn(NO_3)_2$ and 0.25 M $AgNO_3$ to three separate small beakers. After sanding, immerse strips of Cu metal, Zn metal, and Ag metal in each solution, one at a time, for about 30 seconds. Remove the metal strips and note any changes to them or to the solution. Are there any indications of spontaneous reactions? Write the corresponding balanced reactions. Verify with your instructor that you have correctly identified these reactions. Are the observed spontaneous reactions consistent with your knowledge of the activity series? Based on your results, order the three metals in terms of their reactivity (i.e., activity). Verify with your instructor that your conclusions are correct. Sand off any discolored areas on the metal strips after each test.

3. Using the three spontaneous reactions observed in Step 2 as a guide, set up a galvanic cell for each. Use 10.0 mL of the appropriate 0.25 M solution in each half-cell, including the anode half-cell. Record the voltage (emf) as measured by the digital multimeter for each. Using the standard reduction potential table and the Nernst Equation, determine the voltage expected for these galvanic cells and compare to the measured voltage. Is there a consistent deviation between the measured voltages and those predicted from the tables and the Nernst Equation?

    a. Draw a schematic diagram of the galvanic cells, including labels for the direction of electron flow, migration of ions from the salt bridge, and anode and cathode. Simple examples of such diagrams can be found in any general chemistry text.

    b. Verify with your instructor that you have correctly determined the voltages for these galvanic cells and that your diagrams are correct.

4. Set up a galvanic cell involving the Zn/Zn^{2+} half-reaction and the standard hydrogen electrode, by following these steps:

    a. Fill a 250 mL beaker approximately ½ full of 1.0 M HCl. Place the J-shaped platinum electrode in this solution such that the exposed platinum foil is immersed in the acid. Add more acid, if necessary. Why is there no reaction of Pt metal in 1.0 M HCl?

b. Complete the galvanic cell by coupling this beaker with the Zn/Zn^{2+} half-reaction beaker. Measure and record the voltage (emf). What is the half-reaction associated with the acid/Pt half-cell. Use the Nernst Equation to explain why the measured voltage is higher than would be predicted using just the standard reduction potentials for the two half-cells.
c. Wrap a small strip of Mg metal in some copper wire, forming a "handle" with the Cu wire. Using this handle, immerse the Mg metal in the 1.0 M HCl in the beaker, just below the exposed Pt foil on the J electrode. Make sure the bubbles pass over the foil. What are these bubbles, and what is the approximate pressure of the gas in them? Why does the Cu wire not react?
d. Measure and record the voltage (emf) as the gas bubbles pass over the Pt foil. Predict the voltage expected for this cell using the Nernst Equation. Is the difference between the measured and predicted voltage consistent with your earlier observations?
e. Verify with your instructor that your calculations and explanations are correct.

Concentration Effects

5. Make a new galvanic cell, as before, using the Cu/Zn system, using 10.0 mL of the 0.25 M solutions in each half-cell. Measure and record the voltage (emf) of this cell.

   a. Add 10.0 mL of distilled water to the Zn^{2+} solution. Use the Nernst equation to predict the effect that this should have on the voltage of the cell, i.e., determine whether the voltage should increase or decrease. Measure and record the new voltage. Is the measured voltage change consistent with your prediction?
   b. To the Cu^{2+} solution, add a few crystals of solid $CuSO_4$. Use the Nernst Equation to qualitatively predict the effect that this should have on the voltage of the cell. Measure and record the new voltage. Is the measured voltage consistent with your prediction?
   c. Now add a few drops of concentrated ammonia solution to the Cu^{2+} solution. What reaction takes place between Cu^{2+} ions and ammonia? How would this qualitatively affect the voltage of the cell? Measure and record the new voltage. Is the measured voltage consistent with your prediction?
   d. Verify with your instructor that your analysis of these effects is correct.

Concentration Cells

6. Galvanic cells where both half-reactions involve the same materials, e.g., Cu/Cu^{2+}, but where the concentration of the ions are different in the two half-cells, are called "concentration cells." Make a concentration cell using the Cu/Cu^{2+} half-reactions. In one beaker, use 10.0 mL of the 0.25 M $Cu(NO_3)_2$ solution. In the other beaker, add 1 drop of the 0.25 M $Cu(NO_3)_2$ solution to 10.0 mL of distilled water. Assuming that 1 drop of an aqueous solution is 1/20 of a milliliter, what is the concentration of Cu^{2+} ions in the diluted solution? Using the Nernst equation, predict the voltage for such a cell. Measure and record the actual voltage and compare to your prediction. Verify with your instructor that your calculated voltage is correct.

Atmospheric Corrosion

7. In a single, 400 mL beaker about ½ full of distilled water, place Cu and Zn strips, making sure that they are not in contact with each other. Attach the leads through the digital multimeter. Measure the voltage, if any, that is generated by this arrangement. Based on the fact that there are no Cu^{2+} or Zn^{2+} ions in the water to begin with, what would you expect the voltage reading to be? Based on the voltage reading, which metal strip is acting as the anode and which is acting as the cathode? Is there any reaction you can think of that might be going on at the cathode? (Think of all of the species that you know are found in water.)

8. Write the half-reaction that you have thought of in part (a). Add a few drops of phenolphthalein solution to the beaker and wait for about a minute. Is there any evidence to confirm that the half-reaction you have proposed does, in fact, occur? Now switch the multimeter to the appropriate current (amps) setting, and wait for about a minute. Record the measured amperage. Now is there evidence to confirm that the half-reaction you have proposed does, in fact, occur? Shake both of the metal strips to disperse any buildup of chemical products, and switch the multimeter back to the voltage setting to see if the reaction continues. Are these observations consistent with the information about voltage measuring devices given in the Discussion section? Explain.

9. Repeat your observations from Steps 7 and 8, but this time immerse the Cu and Zn metal strips in a solution of 0.50 M NaCl, instead of distilled water. How do your observations differ from those previously made? Verify with your instructor that your analysis of these observations is correct. Save your 0.50 M NaCl solution for the next part.

10. Study the effect of non-uniform dissolved oxygen concentration as follows:

    a. In the large plastic container found in the student drawer under the hood, add 0.50 M NaCl solution until the depth is about ½". Immerse two strips of Zn metal in this solution, on opposite sides of the plastic container, attached to the leads from the digital multimeter. Measure and record the voltage. What voltage would you expect? What are some reasons the measured voltage is not exactly the same as this expected voltage? Now, with a plastic pipet, squeeze bubbles of air near one of the zinc strips. What effect does this have on the voltage? Confirm this effect by squeezing the bubbles of air near the other zinc strip. Explain how this observation is consistent with the explanation of atmospheric corrosion of a ship's hull, sitting in water, referred to in the Discussion section. What are the two half-reactions? Which zinc strip will gradually be eaten away?
    b. Add a few drops of phenolphthalein solution to the solution. Switch the multimeter to the appropriate current (amps) setting, allowing the cell to actively discharge. Continue bubbling air near one of the zinc strips. How does the phenolphthalein help confirm the reactions postulated above?
    c. Verify with your instructor that your analysis of these observations is correct.

Lead Storage and Dry Cell Batteries

11. Make a lead storage battery as follows:
    a. In the instructor's hood, carefully clean two lead strips with sandpaper. Make sure the lead shavings go into the proper waste container, and avoid inhaling the lead dust.
    b. Place the two lead strips on opposite sides of the 400 mL beaker which contains 3.6 M H_2SO_4, found in your hood. Make sure the lead strips do not touch each other.

c. Through the double-throw switch, connect the power supply, the lead strips, and the digital multimeter in such a way that the switch can be set to either include or exclude the external power supply at any given time. Remember, it will be needed during the initial charging and recharging phases of operation, but not during the discharging phase. See the Discussion section for a simple illustration of this type of circuit. Verify with your instructor that your setup is correct.
d. Perform the initial charging of the battery by applying the external power supply for about 5 minutes. Note any reactions occurring during this initial charging process. Are they consistent with the candidate electrolysis reactions mentioned in the Discussion section? If the electrolysis reaction products were trapped in an enclosed space, what possible hazard could this create?
e. Measure the electrical potential generated by this battery. Use the Nernst Equation to predict the expected voltage and compare to that actually measured.
f. Now measure the current produced as this battery discharges. Leave the digital multimeter set to the current scale for just a few seconds. Based on your voltage and amperage values, calculate the power generated by this battery.
g. Insert the light bulb apparatus into the battery circuit. Based on the power calculation, should this particular light bulb light up? (You should be able to see the power requirement label on the light bulb itself.) Test to see if it does. Then remove the light bulb from the battery circuit.
h. Recharge the battery for 2 or 3 minutes.
i. Measure and record the voltage.
j. Allow the battery to discharge for 20 seconds, and measure the voltage again. Repeat every 20 seconds until the voltage drops to nearly zero volts. Make a graph showing the measured voltage vs. time for this battery. (Remember that the battery is only able to discharge when the digital multimeter is set to one of the amperage scales.)
k. Disassemble the lead storage battery, leaving the 400 mL beaker containing the 3.6 M H_2SO_4 in your hood, for use by the next laboratory section.

12. Make a dry cell battery as follows (refer to the paragraphs about dry cells in the Discussion section):
    a. Through the double-throw switch, attach the leads to the zinc plate (anode) and the carbon block (cathode). Include the digital multimeter in the circuit so that the electrical properties of the cell can be measured. See the Discussion section for a simple diagram of such a circuit.
    b. On the wooden block, place the following (in sequence):
        i. The zinc plate
        ii. A piece of filter paper saturated with NH_4Cl solution (the separator)
        iii. A paste consisting of MnO_2 and NH_4Cl (obtain from your instructor; handle this paste with caution, since the MnO_2 stains both hands and clothes)
        iv. The carbon block
    c. Measure and record the voltage generated by this dry cell. Sometimes, the voltage can be increased by improving the contact of the various materials in the cell. This can be done by pressing down on the carbon block with some non-conducting material, e.g., a beaker (caution: pressing too hard will break the carbon block).
    d. Measure the current produced by the cell for a few seconds. Calculate the power generated by the cell. Is there enough power to light the light bulb? If so, test it.
    e. Dispose of the chemicals used to make this cell according to your instructor's directions.

Electrolysis

13. Using the Standard Reduction Potentials table from your textbook, predict the anode and cathode reactions which will occur during the electrolysis of an aqueous solution of KI(aq). Check with your instructor to verify that you have predicted the correct reactions.

14. Set up the electrolytic cell by using the carbon rod as the anode and the platinum J-electrode as the cathode. Include the digital multimeter as part of the circuit so that the current can be measured during the electrolysis. See the discussion section for a diagram of a circuit similar to the one needed here. Test to make sure that the electrodes have been properly identified by electrolyzing a small amount of the 0.20 M KI solution, and verifying that the anode and cathode reaction products are, in fact, produced where expected.

15. Fill a 600 mL beaker about 2/3 full of 0.20 M KI solution. Add the electrodes to the solution. Fill the gas collection tube with 0.20 M KI solution and introduce the inverted tube into the solution without allowing any bubbles to enter the tube. Your instructor may provide some guidance on how to accomplish this. Place the tube over the platinum foil of the J-electrode so that the gas produced there will be trapped in the collection tube. Verify with your instructor that your setup is correct.

16. At a certain marked time, turn on the power supply (at a voltage setting of 9 V) and record the amperage every 30 seconds until the gas collection tube is filled to about the 8 mL mark. Then turn the power supply off.

17. Add a couple of drops of phenolphthalein to the solution. How does this test demonstrate that the reactions occurring during the electrolysis are, in fact, those you predicted?

18. Determine the number of moles of hydrogen gas collected in the tube. Remember to carefully equalize the pressure of the gas with the barometric pressure by adjusting the height of the gas collection tube until the liquid levels inside and outside the tube are equal. Remember to account for the vapor pressure of water in your calculation. Of course, you will need to record the barometric pressure and the temperature!

19. Determine the number of moles of I_2 produced in the electrolysis by titrating the solution with 0.0200 M $Na_2S_2O_3$ solution, using starch near the endpoint. Prior to the titration, acidify the solution by adding about 10 mL of 1.0 M H_2SO_4.

20. Determine the number of moles of electrons which passed from the anode to the cathode during the electrolysis by using the amperage readings recorded over the time of the experiment.

21. Are the calculated moles of hydrogen gas, iodine, and electrons consistent with the anode and cathode half-reactions?

22. If time permits, repeat the electrolysis procedure of steps 15-20. As an optional approach, at the discretion of your instructor, do NOT acidify the solution prior to the thiosulfate titration. In the presence of base, the thiosulfate half-reaction is

$$S_2O_3^{2-} + 6\ OH^- \rightarrow 2\ SO_3^{2-} + 3\ H_2O + 4\ e^- \qquad (9)$$

Show that this titration result is consistent with the results from your first attempt, where the normal thiosulfate reaction applies, i.e.,

$$2\ S_2O_3^{2-} \rightarrow S_4O_6^{2-} + 2\ e^- \qquad (10)$$

23. Discard all of the chemical species down the drain, flushing with plenty of water.

QUESTIONS FOR CONSIDERATION:

1. If Mg metal and 0.25 M $Mg(NO_3)_2$ solutions were coupled in a half-reaction with the Cu/Cu^{2+} system, which metal would serve as the anode, and what would be the expected electric potential generated by the cell?

2. Predict the electric potential generated from a galvanic cell similar to the one built in Step 4, but using 0.010 M HCl instead of 1.0 M HCl.

3. Explain why the atmospheric corrosion was so much more effective in the 0.5 M NaCl compared to when it was carried out in distilled water.

4. Explain why attaching a piece of Zn metal to the hull of a ship helps protect the hull from unwanted corrosion (assume the hull is iron-based). Why would this protection not be effective if the zinc were attached to a painted section of the hull?

5. Write the half-reactions which occur when a lead storage battery is discharging, recharging, and undergoing initial charging.

6. Explain what happens when a "dead" lead storage battery is connected to another battery (e.g., when jump-starting a car with jumper cables). Why is it important to connect the cables to the correct terminals (i.e., electrodes)? Why must the car with the good battery be left running?

7. How would the voltage of a dry cell be affected if the Zn^{2+} ions simply built up near the Zn plate anode?

8. Dry cells are related to alkaline batteries, where the reactants are also Zn(s) and MnO_2(s). However, the electrolyte solution for alkaline batteries is NaOH(aq) instead of NH_4Cl(aq). The products of the half-reactions are ZnO(s) and Mn_2O_3(s), respectively. Write and balance the overall reaction for alkaline batteries.

9. Using the moles of H_2 and I_2 determined in Steps 18 and 19, and the total amount of charge (in coulombs) which passed between the electrodes, determine the value for the Faraday, i.e., the charge of one mole of electrons.

10. Using the results from Millikan's oil drop experiment, i.e., the charge of a single electron is 1.6022×10^{-19} C, and with the result from Problem 9 above, determine the value of Avogadro's number.

11. Determine the total charge transferred during the electrolysis of the KI(aq) solution by applying the Trapezoidal Rule. Compare this result with that obtained by simply averaging all of the current readings. Does the new result improve your calculation of the Faraday or of Avogadro's number (see Problems 9 or 10)?

12. Sketch the cell associated with the atmospheric corrosion of iron. Identify the flow of electrons, the migration of ions, the anode, and the cathode.