Experiment 32: Galvanic Cells

Background Information

- Standard Cell Potential is calculated as $\mathbf{E^o_{cell} = E^o_{red} - E^o_{ox}}$.
- Values for the two half-cell potentials are normally obtained from a table.

Pre-Lab Hints

1. Refer to Figure 32.1 as well as the Chapter 19 class notes (LEO and GER).
   In part c, the metal cations gain electrons to become part of the metal electrode.
2. Refer to the paragraph and accompanying side-notes in the Introduction
   that discuss the salt bridge and the internal circuit (following Figure 32.1).
3. Use $M_1V_1 = M_2V_2$. $M_1 = 0.10$ M. $V_1 = 0.0010$ L. $V_2 = 0.\ 1000$ L.
   The result will have two significant digits.
4. a. The y-intercept is b in equation 32.8. Its value can be found in Figure 32.2.
   b. Find pCu at E = 1.050 V from Figure 32.2. Then, find $[Cu^{+2}]$ using $[Cu^{+2}] = 10^{-pCu}$.
   c. Find pCu using $pCu = -\log[Cu^{+2}]$. Then, find E from Figure 32.2 using that pCu value.
5. a. Reduction (GER) occurs at the cathode (positive half-cell potential).
   b. Oxidation (LEO) occurs at the anode (negative half-cell potential).
   c. First, reverse the anode reaction. Then, apply an integer factor to the cathode reaction.
      Add the two resulting reactions together so that the e^{-1}'s cancel out of the sum completely.
   d. Substitute the two E^O values provided into the equation $E_{cell} = E_{red} - E_{ox}$.
      Do not change any signs or apply any integer factors.
   e. Use the number of cancelled e^{-1}'s from 5c as the n value.
      Write the expression for Q using the overall reaction from 5c.
      Substitute the concentration values provided into the expression for Q to find its value.
      Substitute the value for E_{cell} from 5d, along with the values for n and Q, into equation 32.6.
6. Skip.

Procedure Notes

- Do Part A only. Skip Part B and Part C.
- There are six different cell combinations between the four different half-cells.
- Be sure that each solution is matched with the same metal for its electrode.
- Work in groups of four in all of the steps, except with the unknown.
- Work in groups of two on the unknown.
- Label all of your beakers. Include the solution’s identity as well as its concentration.
- When you are finished with the solutions, place them in the waste jar.
- When you are finished with the metal electrodes, reclean them as in step A1.
  Then, return them to their original locations.
- The black wire connects the left hole of the multimeter to the anode of your cell.
- The red wire connects the right hole of the multimeter to the cathode of your cell.
- Set the multimeter to the 2000 mV setting, and note that the reading is in mV,
  which will need to be converted into V for your report sheet tables.

Lab Questions

1. What is the value of Q at standard state, where all concentrations are 1 M?
   Examine equation 32.6. What is the value of E_{cell} at standard state?

2. When does Q = K? What is the value of E_{cell} in equation 32.6 when Q = K?

3. In this experiment, $Fe_{(s)}$ is used as one of the electrodes.
   However, not every half-cell reaction that involves an iron solution can use $Fe_{(s)}$.
   Show the undesirable cell reaction that will occur between $Fe_{(s)}$ and $Fe^{+3}_{(aq)}$.
   Also, explain how to resolve the issue with the $Fe_{(s)}$ electrode.
   Refer to Table 19.1 (or Appendix E) in your text.

4. Explain why e^{-1}’s flow through the wire from anode to cathode in terms
   of what happens to e^{-1}’s during the oxidation and reduction reactions.
   Which electrode loses e^{-1}’s and which electrode gains e^{-1}’s?

5. The KNO_3 salt bridges are used to prevent the two half-cells from building up charges.
   Explain in terms of charges and e^{-1}’s why each of the two ions (K^{+1} and NO_3^{-1}) migrates
   to its respective electrode (cathode and anode).

Report Sheet for Experiment 32 (Galvanic Cells)

Table 1 E_{cell} Measurements and Half-Cell Reactions

Cell (Anode/ Cathode)	E_{cell} Measured (V)	Anode Metal	Equation for Half-Reaction at Anode	Cathode Metal	Equation for Half-Reaction at Cathode
Zn/Cu					
Mg/Cu					
Fe/Cu					
Mg/Zn					
Mg/Fe					
Zn/Fe					

Table 2 Experimental Reduction Half-Cell Potentials

Cell (Anode/ Cathode)	E_{cell} Measured (V) See Note 1	Cathode Metal See Note 3	Experimental Reduction Half-Cell Potential (V) $E_{Red} = E_{cell} + (-0.76\ V)$	Theoretical Reduction Half-Cell Potential (V)	% Error See Note 2
Zn/Cu		Cu		+ 0.34 V	
Zn/Fe		Fe		− 0.44 V	
Zn/Zn	0.00	Zn		− 0.76 V	
Zn/Mg		Mg		− 2.38 V	
Zn/ Unknown					

Note 1: For Zn/Cu and Zn/Fe, use results from Table 1.
For Zn/Mg, reverse the electrodes used for the Mg/Zn cell.

Note 2: % Error = [(Theoretical minus Experimental) ÷ (Theoretical)] × 100%
Show calculations below.

Note 3: Identify your unknown, which is one of the four metals in the experiment.
Use the theoretical value of that metal in your % error calculation.

1. Write the balanced overall cell reactions for the cells in Table 1.
   Identify the oxidizing agent in each cell.

2. Add the E_{cell} values together for Zn/Cu and Mg/Zn, and compare the sum with the value for Mg/Cu. Add the overall cell reactions (from #1) together for Zn/Cu and Mg/Zn, and compare with the Mg/Cu reaction. What do you observe from these comparisons?

3. Add the E_{cell} values together for Zn/Fe and Mg/Zn, and compare the sum with the value for Mg/Fe. Add the overall cell reactions (from #1) together for Zn/Fe and Mg/Zn, and compare with the Mg/Fe reaction. What do you observe from these comparisons?