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**Date:** 4/21/2021

**Instructor:** Guamani Gyanwai

**Experiment #11:** Electrochemistry: Galvanic Cells and Electroplating

**Objective:**

- To set up the half cells using proper electrodes and solutions.
- To calculate standard cell potential for galvanic cells made of those half cells.
- To measure the standard cell potentials for the galvanic cells.
- To electroplate a metal using another metal using an electrolytic cell

**Materials:** Al Cu.

**Part A:**

- Pure metal electrode: Zn, Cu, Al, Fe, Mg, Ni
- Six 100 ml beakers
- One 150 ml beaker
- 1 M $Zn(NO_3)_2$
- 1 M $CuSO_4$
- 1 M $Al(NO_3)_3$
- 1 M $FeSO_4$
- 1 M $Mg(NO_3)_2$
- 1 M $NiSO_4$
- Vernier charge sensor
- Computer interface
- Wire with clips
- 1M $KNO_3$

**Part B:**

- 250 ml beaker
- Copper electrodes
- Aluminum foil
- 1 M $CuSO_4$
- Wire with clips
- 1.5V dry cell

**Procedure:**

- Part A: Six 100 ml beakers were brought and labeled with the appropriate metal symbol for each metal provided. Then the six salt solutions were brought and 20 ml of each of the salts were put into the respective metal beakers. Then each metal rod was placed in each respective solution. Then in a 150 ml beaker about 20 ml of $KNO_3$ was poured into the beaker for the salt bridge solution. Next the salt bridge was made using a paper towel dipped in the $KNO_3$ solution. Then using the planning table in the instruction page, the combination of half cells were found. Then using the cell 1 on the planning chart the positive wire from the volt meter was placed to the cathode Fe meatal rod and the negative to the anode Zn metal rod. After the voltage stabilized it was recorded on the table. Then the same steps were took with the rest of the combination of half cells.
- Part B: A 250 ml beaker was brought and filled about 1/3 full of 1M $CuSO_4$ solution. Then a copper electrode was brought and attached to the positive terminal of a battery with the wire and clip. Then a piece of aluminum foil about the same size as the copper was attached to the negative side the clip. Next the electrodes were dipped in the copper sulfate solution for 5 minutes. And the observations were recorded.

**Observation:**

Part A:

Part A: Galvanic Cells

Provided metal electrodes: Pb, Zn, Al, Cu, Fe

Provided salts: $CuSO_4$, $AlCl_3$, $FeSO_4$, $Zn(NO_3)_2$, $Pb(NO_3)_2$

Measured $E_{cell}$ Voltage Table

| $E_{cell}$ (Volts) | $Zn/Zn^{2+}$ | $Fe/Fe^{2+}$ | $Al/Al^{3+}$ | $Cu/Cu^{2+}$ | ~~$Ni/Ni^{2+}$~~ $Pb/Pb^{2+}$ | ~~$Mg/Mg^{2+}$~~ |
|---|---|---|---|---|---|---|
| $Zn/Zn^{2+}$ | --- | 1) 0.06 | 0.06 | 2) 0.3 | 3) 0.12 | |
| $Fe/Fe^{2+}$ | --- | --- | 0.14 | 4) 0.26 | 5) 0.08 | |
| $Al/Al^{3+}$ | --- | --- | --- | 6) 0.38 | 0.2 | |
| $Cu/Cu^{2+}$ | --- | --- | --- | --- | 0.18 | |
| ~~$Ni/Ni^{2+}$~~ $Pb/Pb^{2+}$ | --- | --- | --- | --- | --- | |
| ~~$Mg/Mg^{2+}$~~ | --- | --- | --- | --- | --- | |

Now choose ONLY SIX cells out of the total fifteen measured $E_{cell}$ you want to calculate the $E^{\circ}_{cell}$ from above combinations. Mark those cells on the table above for your simplicity. Write Half-cell reactions.

**Cell 1:** Anodic half-cell reaction $Zn \rightarrow Zn^{2+} + 2e^-$ +0.76

Cathodic half-cell reaction $Fe^{2+} + 2e^- \rightarrow Fe$ -0.44

**Cell 2:** Anodic half-cell reaction $Zn \rightarrow Zn^{2+} + 2e^-$ +0.76

Cathodic half-cell reaction $Cu^{2+} + 2e \rightarrow Cu$ +0.34

**Cell 3:** Anodic half-cell reaction $Zn \rightarrow Zn^{2+} + 2e^-$ +0.76

Cathodic half-cell reaction $Pb^{2+} + 2e^- \rightarrow Pb$ -0.13

**Cell 4:** Anodic half-cell reaction $Fe \rightarrow Fe^{2+} + 2e^-$ +0.44

Cathodic half-cell reaction $Cu^{2+} + 2e^- \rightarrow Cu$ +0.34

**Cell 5:** Anodic half-cell reaction $Fe \rightarrow Fe^{2+} + 2e^-$ +.44

Cathodic half-cell reaction $Pb^{2+} + 2e^- \rightarrow Pb$ -0.13

**Cell 6:** Anodic half-cell reaction $Al \rightarrow Al^{3+} + 3e^-$ +1.66

Cathodic half-cell reaction $Cu^{2+} + 2e^- \rightarrow Cu$ +0.34

Calculate the $E^{\circ}_{cell}$ using above reactions as your guide to find the values in Table 11.2 in the Introduction section and using Equation (iii).

| Cell Number | $E^{\circ}_{cell}$ (Calculated) | $E^{\circ}_{cell}$ (Measured) | Difference or % Error | Comments |
|---|---|---|---|---|
| 1 | 0.32 | 0.06 | 81% | |
| 2 | 1.1 | 0.3 | 73% | |
| 3 | 0.63 | 0.12 | 81% | |
| 4 | 0.78 | 0.26 | 67% | |
| 5 | 0.31 | 0.08 | 74% | |
| 6 | 2 | 0.38 | 81% | |

**Part B:** The Cu was used as the cathode and the Al was the anode. The cathode was positive and the anode was negative. After 5 minutes the copper was dissolving and the Aluminum was turning into copper.

**Discussion:**

- In part A the voltage was lower in our experiment then of the other students. And in part B the cathode was giving its properties to the anode.

**Conclusion:**

- First half cells using proper electrodes and solutions were set up. Next standard cell potential for galvanic cells made of those half cells were calculated and came out to be ( from cell 1 to cell 6 respectively) 0.32, 1.1, 0.63, 0.78, 0.31, and 2. Then the standard cell potentials for the galvanic cells were measured. Finally a metal was electroplated using another metal using an electrolytic cell.

## Materials and Chemicals

**Part A:** Pure metal electrodes: Zn, Cu, Al, Fe, Mg, Ni, six 100-mL beakers, one 150-mL beaker, salt solutions: 1M each of $Zn(NO_3)_2$, $CuSO_4$, $Al(NO_3)_3$, $FeSO_4$, $Mg(NO_3)_2$, $NiSO_4$, Vernier charge sensor, computer interface, wire with clips, 1.0M $KNO_3$ solution
**Part B:** 250-mL beaker, copper electrode, aluminum foil, 1.0 M $CuSO_4$ solution, wire with clips, 1.5V dry cell.

## Procedure

### Part A: Galvanic Cells

1. Bring six 100-mL beakers, wash them well and dry with a paper towel. Label them with the appropriate metal symbol. For example: the beaker used for $Fe/Fe^{2+}$ half-cell should be labelled as *Fe*.
2. Bring the provided six salt solutions and pour 20 mL of the salt solution into the respective metal beakers.
3. Connect the Vernier charge sensor to the computer interface. Set the charge sensor to the minimum voltage setting available. You can change this setting when needed for higher voltage (if any).
4. The interface display will show the micro ampere (μA) or milliampere (mA) of the current measured. Tap on the numbers displayed, you will be given option to change the units, choose the unit Volts.
5. Place each metal rod into the respective solution. These are your six half-cells. You will make total 15 cells using combinations of these half-cells.
6. In the 150-mL beaker, pour about 20mL 1.0M $KNO_3$ solution. This is your salt bridge solution. To make a salt bridge, you will need a paper towel of correct size. Dip it in the KNO3 solution. The wet paper towel is your salt bridge.
7. Before starting to construct the cells using the half electrodes, plan well how you want to proceed with the combination of half-cells. Consider the following:
   A. Checking the Standard Reduction Potential Table, decide which is cathode and which is anode. This will help you with terminal connection, (+) wire or (-) wire?
   B. Make a good plan how you want to proceed with half-cell combinations. Which half cell to take first and proceed further. There are fifteen different combinations possible out of six half-cells. The first one is done for you.
   C. Each time you make a cell using the two half-cells, use a new salt bridge. You will need 15 salt bridges for 15 cells.

**Planning Table: Cell 1 is already planned for you.**

| Cell # | Combination | Anode (−) | Cathode (+) |
|---|---|---|---|
| Cell 1 | Zn-Fe | Zn | Fe |
| Cell 2 | Zn-Al | Al | Zn |
| Cell 3 | Zn-Cu | Zn | Cu |
| Cell 4 | Zn-Pb | Zn | Pb |
| Cell 5 | Al-Fe | Al | Fe |
| Cell 5 | Cu-Fe | Fe | Cu |
| Cell 5 | Pb-Fe | Fe | Pb |
| Cell 5 | Cu-Al | Al | Cu |
| Cell 9 | Pb-Al | Al | Pb |

| | | | |
|---|---|---|---|
| Cell 10 | Pb-Cu | Pb | Cu |
| Cell 11 | | | |
| Cell 12 | | | |
| Cell 13 | | | |
| Cell 14 | | | |
| Cell 15 | | | |

8. Cell 1: Attach the positive (+) wire from the charge sensor to the cathode = Fe (metal rod) with the clip. Attach the negative (-) wire from the charge sensor to the anode = Zn metal rod with the clip. Connect the two half-cell's solutions with a salt bridge. Your voltage of the cell will be displayed on your interface. Let it stabilize and record the value.
9. Repeat this for all the rest of the combinations you chose without any mix-up or confusion. Make sure the half-cells and the salt bridge should not be contaminated. If there is a contamination, you may need to start over or may remake the half-cell and proceed.
10. Once finished, take the electrodes out, wash them with water well and wipe with dry paper towel.
11. Dispose of the solutions as directed and wash the beakers with soap and warm water.

**Part B: Electroplating**

1. Bring a clean 250-mL beaker. Add the provided 1.0 M $CuSO_4$ solution about 1/3rd full.
2. Bring a copper electrode and attach to the positive terminal of a battery with the wire with the clip.
3. Bring an aluminum foil piece about the same size as the copper electrode. Attach the foil with the clip to the negative terminal of the battery.
4. Dip the two electrodes into the copper sulfate solution of the beaker. Let it sit for 5 minutes.
5. You will start to see changes to the aluminum foil as well as the copper electrode. Record your observation on the data page.

Name Brandon Deramus

## Experiment 11 | Electrochemistry: Galvanic Cells and Electroplating

Pre-laboratory Questions

*Complete this page before coming to the lab. Read the information given to answer the following questions.*

Q1. What type of reactions do electrochemical process use?

redox reactions.

Q2. What is the main difference between the reactions in a galvanic cell and those in electrolytic cells?

galvanic is spontaneous and electrolytic is non-spontaneous

Q3. What is a half-cell? How do we devise a half cell for a Zn electrode?

metal rod dipped in a salt solution of the same metal. $ZnSO_4$

Q4. Take the reference from Fig. 11.2 and Table 11.2 to decide which metal is anode and which is cathode while making a galvanic cell between $Fe/Fe^{2+}$ and $Al/Al^{3+}$.

$Al/Al^{3+}$ is anode and $Fe/Fe^{2+}$ is cathode

Q5. How do we know which metal becomes an anode and which a cathode in an electrolytic cell?

The metal with the lower $E^0$ will be an anode and higher will be cathode.

***Instructor's approval*** GG

$Zn(NO_3)_2$

1M in 250 ml

$$\frac{297.49\,g}{1\,mol} \quad \frac{1\,mol}{250}$$

$$\frac{mol}{.250\,mol} = 1$$

$$.25\,mol = \frac{g}{297.49\,g} = 74.37$$

Date Brandon Deramus

Name 4/14

Lab Partner Abby Nance

## Experiment 11 | Electrochemistry: Galvanic Cells and Electroplating

Data Sheet

**Part A: Galvanic Cells**

Provided metal electrodes: Pb, Zn, Al, Cu, Fe

Provided salts $CuSO_4$, $AlCl_3$, $FeSO_4$, $Zn(NO_3)_2$, $Pb(NO_3)_2$

Measured $E_{cell}$ Voltage Table

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|---|---|---|---|---|---|---|
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| $Fe/Fe^{2+}$ | --- | --- | 0.14 | 4 0.26 | 5 0.08 | |
| $Al/Al^{3+}$ | --- | --- | --- | 6 0.38 | 0.2 | |
| $Cu/Cu^{2+}$ | --- | --- | --- | --- | 0.18 | |
| ~~$Ni/Ni^{2+}$~~ $Pb/Pb^{2+}$ | --- | --- | --- | --- | --- | |
| ~~$Mg/Mg^{2+}$~~ | --- | --- | --- | --- | --- | --- |

Now choose ONLY SIX cells out of the total fifteen measured $E_{cell}$ you want to calculate the $E^o_{cell}$ from above combinations. Mark those cells on the table above for your simplicity. Write Half-cell reactions.

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Cathodic half-cell reaction $Cu^{2+} + 2e^- \rightarrow Cu$ +0.34

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**Part B: Electroplating**

Metal used as cathode Cu

Metal used as anode Al

Terminal of the battery cathode was connected to +

Terminal of the battery anode was connected to —

Observation of the cathode after 5 minutes Copper was coming off

Observation of the anode after 5 minutes Turning into copper

Brief comment on the observations The cathode was giving to the anode.