EXPERIMENT/SUBJECT	DATE	08
LAB PARTNER [redacted]	LOCKER/DESK NO.	COURSE & SECTION NO.

Purpose: To create simple galvanic cells to measure and compare actual cell potentials to theoretical cell potentials.

Procedures:

PART ONE

Zn(s) / Zn^{2+} (0.05M)

Cu^{2+} (0.05M) / Cu(s)

$Zn(s) + Cu^{2+}_{(aq)} \rightarrow Zn^{2+}_{(aq)} + Cu(s)$

1) Clean a strip of each metal, Cu(s) and Zn(s) with emery paper.
2) Clean and rinse two 50mL beakers with water.
3) Fill one beaker with 20mL of copper sulfate and the other with 20mL of zinc sulfate.
4) Connect the two beakers with a salt bridge which is a 1.0 M KNO_3 solution in water.
5) Place the copper in the copper sulfate solution and the zinc metal in the zinc sulfate solution.
6) Complete the circuit by connecting the voltmeter ~~to~~.
7) Record the reading on the voltmeter.

CELL DIAGRAM

Zn(s) | Zn^{2+}(aq 0.05M) || Cu^{2+}(aq 0.05M) | Cu(s)

PART TWO

- Refer to question # 2 in Prelab.

CELL DIAGRAM

Cu^{2+}(aq 0.05M) | Cu(s) || Cu^{2+}(0.00005M) | Cu(s)

PART THREE

(1) Obtain an Ecell value from the TA and compare this value to the standard potential values in prelab question # 1.

(2) use the cell potential that is the closest higher value than the value obtained from the TA. ~~since our value was 0.97, we found that~~

(3) since the electrochemical cell including the oxidizing agent Cu^{2+} and reducing agent Zn^{2+} has a value of 1.03V, which is closest to and greater than the obtained value of 0.97, use this value.

(4) Use the nerst equation to solve for Q, using Ecell and the standard potential as follows:

$0.97V = 1.103V - (0.0592/2)\log Q$

$E_{cell} = E°_{cell} - \frac{0.0592}{2}\log Q$

$\log Q = \frac{2 \times (0.97 - 1.103)}{-0.0592}$

$= 4.4932$

$\therefore Q = 31134S.9661$

$= 0.089$

(5) since the highest value of Zn^{2+} is 0.05M, and Q = $[Zn^{2+}_{aq}] / [Cu^{2+}_{(aq)}]$, solve for the concentration of Q.

$311345.9661 = \frac{0.05M}{[Cu^{2+}_{(aq)}]}$

$[Cu^{2+}] = 1.6 \times 10^{-6} M$.

SIGNATURE	DATE	WITNESS/TA	DATE