EXPERIMENT/SUBJECT	DATE

LAB PARTNER	LOCKER/DESK NO.	COURSE & SECTION NO.

Because the concentration of $Cu^{2+}(aq)$ is very small, we rounded this value to 2.5×10^{-6} M.

(6) Dilution of Cu^{2+} to 2.5×10^{-6} M from an original concentration of 0.05 M proceeds as follows.

a) solve for V2, if 0.5 mL of 0.05 M $Cu^{2+}(aq)$ must be diluted to 5×10^{-4} M.

$C_1V_1 = C_2V_2$

$(0.05M)(0.5mL) = (5 \times 10^{-4} M)(V_2)$

$V_2 = 50 mL.$

b) Next the 5×10^{-4} M $Cu^{2+}(aq)$ solution must be diluted further to 5×10^{-6} M. Using 0.5 mL of this solution, the amount of water needed was calculated as follows:

$C_1V_1 = C_2V_2$

$(5 \times 10^{-4} M)(0.5mL) = (5 \times 10^{-6} M)(V_2)$

$V_2 = 50 mL.$

c) Now the 5×10^{-6} M solution must be diluted to the final concentration of 2.5×10^{-6} M. Using 20 mL of the solution, the amount of water needed was calculated as follows:

$C_1V_1 = C_2V_2$

$(5 \times 10^{-6} M)(20 mL) = (2.5 \times 10^{-6} M)(V_2)$

$V_2 = 40 mL$

Thus the final concentration was 40 ~~mL~~ $Cu^{2+}(aq)$.

(7) obtain 40 mL of $Zn^{2+}(aq)$ solution in a clean beaker.

(8) clean Cu(s) and Zn(s) with emery paper and insert into respective solutions.

(9) connect each metal to the voltmeter.

(10) insert a salt bridge and record potential on voltmeter.

Cell Diagram

$Zn(s)|Zn^{2+}(aq, 0.05M)||Cu^{2+}(aq, 2.5 \times 10^{-6} M)|Cu(s)$

CALCULATIONS (% ERROR)

PART ONE

$\% \text{ ERROR} = \frac{(\text{actual} - \text{theoretical})}{\text{theoretical}} \times 100$

$= \frac{1.02V - 1.039V}{1.039V} \times 100\%$

$= -7.52\%$

PART TWO:

$E_{cell} \% \text{ ERROR} = \frac{(\text{actual} - \text{theoretical})}{\text{theoretical}} \times 100\%$

SIGNATURE	DATE	WITNESS/TA	DATE