# PREPARATION WORKSHEET

Name: Alex Koch Date: ____________

TA: ____________ Section: ____________

*Answer the following questions using information obtained from lecture, the textbook, or lab manual. The responses will be collected at the* ***beginning of recitation*** *and checked at the start of your lab section for credit.*

1. The standard reduction potentials for $Cu^{2+}/Cu$ and $Zn^{2+}/Zn$ are given below:

   **$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$   +0.34 V**

   **$Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$   -0.76 V**

   a. What is the equation for determining the standard potential for an electrochemical cell?

   $$E_{cell} = E^\circ_{cell} - \frac{RT \ln(Q)}{nF}$$

   b. What is the equation that determines if an electrochemical will react spontaneously?

   $$\Delta G = \Delta G^\circ + RT \ln(Q)$$

   c. If a strip of copper metal is placed into a solution of $ZnSO_4(aq)$, will a reaction occur? Explain why and write the net ionic equation for the reaction, if any.

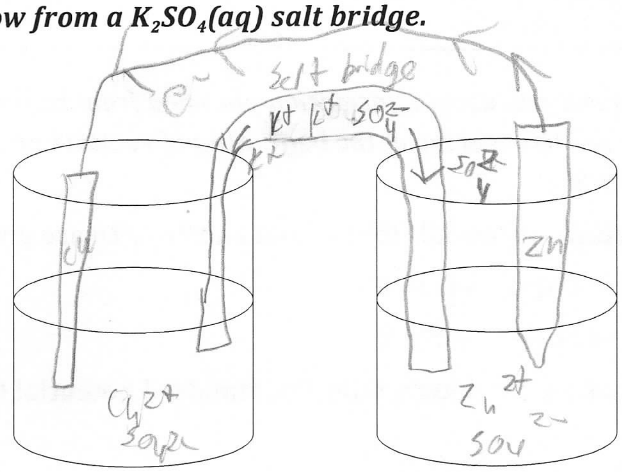
   No, zinc ~~is more reactive~~ than copper has a ~~higher~~ lower ~~reactio~~ reduction potential

   d. If a strip of zinc metal is placed into a solution of $CuSO_4(aq)$, will a reaction occur? Explain why and write the net ionic equation for the reaction, if any.

   Yes, zinc ~~is more reactive~~ has a lower reduction potential

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

**e.** Complete the voltaic cell set-up below assuming it is a Cu–Zn cell utilizing Cu and Zn solid strips, and $CuSO_4(aq)$ and $ZnSO_4(aq)$ solutions in the beakers. Draw and label all missing components. ***Include the direction of electron flow and the direction at which ions flow from a $K_2SO_4(aq)$ salt bridge.***




**2.** Consider an electrochemical cell similar to Figure 10.1 with a Pb(*s*) electrode and $Pb(NO_3)_2(aq)$ solution in one beaker, and Ag(*s*) and $AgNO_3(aq)$ in the other. The aqueous solutions both have initial concentrations of 1.0 M.

**$Ag^+(aq) + e^- \rightarrow Ag(s)$    +0.80 V**

**$Pb^{2+}(aq) + 2e^- \rightarrow Pb(s)$    -0.26 V**

**a.** Write down the half-reactions, and derive the overall balanced reaction. Also, show the calculation for the standard cell potential.

$Ag + OH^- \rightarrow AgO^- + e^- + H^+$

$PbO + H_2O + 2e^- \rightarrow Pb + 2OH^-$

$2Ag + PbO + H_2O \rightarrow AgO^- + Pb + H^+$

$0.80 - (-0.26) = 1.06\ V$

**b.** As time passes, how will the mass of the Pb electrode change?

increase

**c.** Referring to the Nernst Equation, if the $Pb^{2+}(aq)$ concentration is kept at 1.0 M, but the $Ag^{+}(aq)$ concentration is increased to a much higher concentration, how will the cell potential change compared to the standard cell potential?

Increases

moles (n) is in the denominator of $-\frac{RT \ln(a)}{nF}$

**d.** If both the $Pb^{2+}(aq)$ and $Ag^{+}(aq)$ concentrations are changed to 0.01 M, how will the cell potential change, compared to the standard cell potential?

stays the same

**3.** Write down the Nernst equation for the cell reaction described in the previous problem. Show in detail the calculation of the cell potential in the case where $Pb^{2+}(aq)$ and $Ag^{+}(aq)$ concentrations are both 0.01 M. Check that your prediction in the previous question agrees with your answer here.

1M | 0.01 M

Ag $1.06 - \frac{0.0592}{2} \log(1) = 1.06$ | $1.06 - \frac{0.0592}{0.02} \log\left(\frac{0.01}{0.01}\right) = 1.06$

Pb $1.06 - \frac{0.0592}{2} \log(1) = 1,06$ | $1.06 - \frac{0.0529}{0.02} \log(1) = 1.06$

**4.** Part D of this experiment examines the effect of varying the concentrations of solutions in the anode and cathode compartments. Concentrations of 1.0 M, $1.0 \times 10^{-2}$ M, $1.0 \times 10^{-4}$ M will be examined. While the 1.0 M concentrations are readily available, the other two concentrations should be prepared. For each of the desired solutions on the table below, calculate the amount of DI water to be added to the given stock solution.

**TABLE 10.1**

| | Desired Solution | Stock Solution | DI Water to Add to Stock Solution (mL) |
|---|---|---|---|
| 1 | $1.0 \times 10^{-2}$ M $CuSO_4$ | 1 mL of 1.0 M $CuSO_4$ | 99 |
| 2. | $1.0 \times 10^{-4}$ M $CuSO_4$ | 1 mL of $1.0 \times 10^{-2}$ M $CuSO_4$ | 99 |
| 3. | $1.0 \times 10^{-2}$ M $ZnSO_4$ | 1 mL of 1.0 M $ZnSO_4$ | 99 ml |
| 4. | $1.0 \times 10^{-4}$ M $ZnSO_4$ | 1 mL of $1.0 \times 10^{-2}$ M $ZnSO_4$ | 99 |

**5.** Show the calculations for determining the volume of DI water to add to the given stock solution for the preparation of $1.0 \times 10^{-2}$ M $CuSO_4$ and $1.0 \times 10^{-4}$ M $CuSO_4$ solutions.

1. $1\,mL \cdot 1\,M = 1 \cdot 10^{-2}\,M \cdot X$

   $X = 100\,mL$ $\quad$ $100\,mL - 1\,mL = 99\,mL$

2. $1\,mL \cdot 1 \cdot 10^{-2}\,M = 1 \cdot 10^{-4}\,M \cdot X$

   $X = 100\,mL$ $\quad$ $100\,mL - 1\,mL = 99\,mL$

3. $1\,mL \cdot 1\,M = 1 \cdot 10^{-2}\,M \cdot X$

   $X = 100\,mL$ $\quad$ $100\,mL - 1\,mL = 99\,mL$

4. $1\,mL \cdot 1E{-2} = 1 \cdot 10^{-4} \cdot X$

   $X = 100\,mL$ $\quad$ $100\,mL - 1\,mL = 99\,mL$

   total vol