Experiment # 2

Title of experiment:

Concentration Cell- Nernst equation

Aim:

To demonstrate the concentration-dependence of the potentials of the half cells with different electrolyte concentrations

Introduction:

The magnitude of the potential difference in Nernst equation depends on the potential difference between metals, concentration and temperature. The general relationship is given by the following equation (1)

$$E_{cell} = E^{o}_{cell} - \left(\frac{RT}{nF}\right)(\ln Q) \quad (1)$$

If the anode and the cathode are the same metal, we will have a concentration cell.

A concentration cell acts to dilute the more concentrated solution and concentrate the more dilute solution, creating a voltage as the cell reaches an equilibrium. This is achieved by transferring the electrons from the cell with the lower concentration to the cell with the higher concentration.

The Nernst equation is given by:

$$E = E^{\circ} - \frac{0.0257}{n} \ln \left(\frac{[Cu^{2+}]_{anode}}{[Cu^{2+}]_{cathode}} \right) \quad (2)$$

$E^{\circ}_{cell} = E^{\circ}_{cathode} - E^{\circ}_{anode}$ = zero

$$E = -\frac{-0.059}{2} \lg \left[\frac{[Cu+2]anode}{[Cu+2]cathode} \right] \quad (3)$$

In this experiment we will demonstrate the dependence of the potential on concentration by using two different concentrations cupper metal. The reactions in anode and cathode are:

Less concentrated is anode $Cu(s) \rightarrow Cu^{2+}(aq)A + 2e^{-}$
More concentrated is cathode $Cu^{2+}(aq)C + 2e^{-} \rightarrow Cu(s)$

The net reaction is $Cu^{2+}(aq)C \rightarrow Cu^{2+}(aq)A$

The difference in concentration leads to a concentration gradient and produces a voltage difference and a current flow where oxidation occurs in the anode to increase the concentration of Cu^{2+} in the solution and reduction occurs in the cathode to decrease the concentration of Cu^{2+}. The voltage decreases until the Cu^{2+} concentration of both sides is equal and the voltage approaches zero.

Materials:

Volt meter (pH meter)- 50 mL Beakers (2) -two connection leads with Alligator clips - Sand paper

-Cu strips (or sheets)- Solutions of Copper sulfate 1M- 0.010 [Cu^{2+}] - salt bridge filled with 3M KNO3

Procedure:

1. Each pair of student will get 2 sheets of Cu. The Cu sheets are washed with distilled water then dried.
2. Fill a clean 50-mL beaker with 30 mL of 1 M $CuSO_4$ solution. Connect the alligator lead to the Cu Sheet and immerse half of the sheet into the solution. Plug the lead into the positive side of the Voltmeter, this is your cathode half-cell.

2.Prepare the anode Cu half cell by immersing the Cu sheet into a beaker filled with 30 mL solution of [Cu^{2+}] = 0. 1M. Connect the alligator clip to the copper sheet in the beaker and plug the lead into the negative side of the Voltmeter.

3. Be sure the dial on the voltmeter is set to DC Volts and after 30 seconds record the voltage. Does this voltage reading make sense? Record the voltage every 10 seconds until it reaches equilibrium.

4. Put the Voltmeter on hold. Dispose of the 1 M solution in the cathode half -cell and fill a clean beaker with 30 mL of 0.010 M $CuSO_4$ solution. Add a fresh solution in the anode half –cell 0.001 M. Turn on the Voltmeter. After about 10 seconds, read the voltage.
5. Record the results in the table below.
6. Put the Voltmeter on hold after the voltage reach equilibrium. Add 20 mL of distilled water in the cathode half –cell. Read the voltage. What do you notice??

Results:

Report your data in the table below

$CuSO_4$ solution	Voltage	Anode solution	Cathode solution
0.10M/1 M			
0.001M/0.01M			
Dilution of cathode solution			

Report:

1. From your observations, for every combination of difference electrolytes concentrations , does the potential difference decreases with time? Why?
2. Compare the measured & theoretical values for these cell potentials.

From your experiment, result which deviate greatly from the theoretical

value may arise when more concentrated (1.0M/0.1M) OR when highly diluted (<0.001M) solutions are used. Give your reasons.

3. What is your conclusion ?