**Exploring Electrochemistry**

**PURPOSE**

* Create a galvanic (Galvanic) cell and determine the voltage generated by it
* Use the Nernst equation to calculate the concentration of an unknown solution in a galvanic cell
* Create a concentration cell and use the Nernst equation to calculate the concentration of an unknown solution
* Predict an activity series

**INTRODUCTION**

The **electromotive force** (“emf”) of a galvanic cell may be calculated by combining the values for each half-cell from a table of Standard Reduction Potentials to find E°, the **standard voltage** for the cell. The superscript (“°”) indicates standard conditions of 1 ATM, 298 K, and 1 M concentration of all species in solution. If any of these characteristics differs from the standard condition, then the voltage in the cell will be changed. To calculate the altered or nonstandard voltage, you can use the Nernst equation, developed by the Nobel laureate Walther Nernst. In the full form of the Nernst equation, a difference in any of these characteristics can be accommodated. This form is

|  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- |
|  |  |  |  | ( | RT | ) |  |
| E | = | E° | - | ------ | lnQ |
|  |  |  |  | nF |  |

and is often expressed in terms of common (base 10) logarithms in the form

|  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- |
|  |  |  |  | ( | 2.303RT | ) |  |
| E | = | E° | - | ------------ | logQ |
|  |  |  |  | nF |  |

where E = nonstandard voltage; R = 8.314 J mol-1 K-1; T = temperature in Kelvins; n = number of moles of electrons transferred in the redox; F = Faraday’s constant (96,500 coulombs mol-1); and Q = trial equilibrium reaction quotient, expressed as

|  |
| --- |
| [products] |
| ------------ |
| [reactants] |

But wait! There is an even simpler version of the equation. Because ions in solution are not affected by pressure over the surface of the solution, and one frequently assumes the standard temperature of 298 K, there is a form of the equation that allows calculation with only a difference in concentrations of ions of the two half-cells.

|  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- |
|  |  |  |  | ( | 0.0591 | ) |  |
| E | = | E° | - | ------------ | logQ |
|  |  |  |  | n |  |

Because the cell potential depends on the concentration of the ions in the two half-cells, a galvanic cell can be constructed with identical half-cells except for the concentrations of the ions. This is called a **concentration cell**. The difference in voltage is apt to be small because all components of the half-cells are identical except for the concentrations. For example, the voltage between a 1.0 M Cu2+ solution half-cell and one of 0.010 M Cu2+ solution is just less than 0.06 V.

Experimentation with various metals and solutions containing the cations of those metals is fascinating. Observation of such situations can be used to establish an **activity series** of the ease of oxidation of certain metals. For example, if you place a piece of sold copper wire into a solution of silver nitrate, solid silver metal can be observed forming onto the copper wire (thus the reduction of Ag+ to Ag), while the solution in the container becomes blue (thus Cu becomes Cu2+). From this, it can be concluded that copper is more easily oxidized than silver. This conclusion can be verified by dropping a piece of silver wire into a solution of Cu2+ and noticing that nothing happens, because the more easily oxidized copper is already oxidized.

In this experiment, microscale techniques will be used to set up three different electrochemical cells and thus compare the behavior of three different half-cells paired with each other.

**Pre-Lab Questions**

1. What is a Galvanic cell?
2. In a functioning galvanic cell
3. at which electrode does oxidation occur? reduction occur?
4. which electrode gains mass? loses mass?
5. around which electrode does the concentration of cation solution increase? decrease?
6. from which electrode and to which electrode do electrons in the external circuit move?
   1. What is the function of a salt bridge?
   2. A standard cell will be created using Zn/Zn+2 and Cu+2/Cu as the electrodes
      1. Write the balanced equation for the reaction that occurs in this galvanic cell.
      2. Calculate the overall voltage for this cell.
      3. Which electrode is at the anode? Which electrode is at the cathode?
   3. Has all chemistry ceased when the voltage of a galvanic cell becomes zero? Explain.
   4. Write the equation used to find the voltage of a non-standard cell at 25 C.

**MATERIALS**

24-well microplate Cu metal strip

1 M CuSO4 Zn metal strip

1 M ZnSO4 Filter paper

Unknown M CuSO4 Forceps

1 M KCl Voltmeter (or other voltage data collection device)

Sandpaper (or steel wool)

**PROCEDURE**

**1. Initial Preparations**

1. Cut a piece of filter paper into a number of strips approximately 0.5 by 4.0 cm.
2. Place the strips of filter paper into a small beaker containing 1 M KCl solution, allowing them to soak up the solution.
3. Locate two strips of copper metal and one strip of zinc metal, approximately 0.5 by 4.0 cm each.
4. Use fine sandpaper or steel wool to polish the metal strips. Try to clean off any external corrosion to leave a shiny metal surface. Carefully wipe away any loose particles from the surface of the strips.
5. Select three clean wells in a 24-well microplate. The wells should be adjacent to each other, in a triangle.
6. Fill one well about ¾ full with 1 M CuSO4 solution. Repeat for the other two wells with 1 M ZnSO4 and the unknown copper (II) solutions.
7. Place a strip of Cu metal into each of the two wells containing Cu2+ solution and a strip of Zn metal into the well containing the Zn2+ solution. Try to stand each strip at the out edge of its microwell, such that an equilateral triangle is formed with the metal strips at the apexes of the triangle.

**2. Creating a Galvanic Cell**

1. Use the forceps to retrieve one filter paper strip from the 1 M KCl solution. Allow excess solution to drip from the strip back into the container.
2. Use the forceps to drape the paper strip from the well containing 1 M CuSO4 solution, across the top of the microwell plate, and down into the well containing the 1 M ZnSO­4 solution. This paper acts as the salt bridge between the two half-cells. Do not allow the paper to touch the metal strips.
3. Quickly touch the leads of the voltmeter to the copper strip in one well and to the zinc strip in the adjacent well. Try not to put too much pressure on the metal strips. This will cause them to fall and possibly touch the salt bridge. The best place to touch the strips is about half way down, being careful not to immerse the lead in the solution.
4. Record the voltage measured by the voltmeter. If the voltage reading is negative, switch the leads with the metals to see a positive reading. Since all of the cell reactions in this experiment are spontaneous, all voltage readings should be positive.
5. Remove the leads from the metal strips.
6. Use the forceps to retrieve the salt bridge and dispose of it in the trash.

**3. Creating a Concentration Cell**

1. Use the forceps to retrieve one filter paper strip from the 1 M KCl solution. Allow excess solution to drip from the strip back into the container.
2. Use the forceps to drape the paper strip from the well containing 1 M CuSO4 solution, across the top of the microwell plate, and down into the well containing the unknown concentration CuSO4 solution. This paper acts as the salt bridge between the two half-cells. Do not allow the paper to contact the metal strips.
3. Quickly touch the leads of the voltmeter to the copper strip in one well and to the copper strip in the adjacent well. Try not to put too much pressure on the metal strips. This will cause them to fall and possibly touch the salt bridge. The best place to touch the strips is about half way down, being careful not to immerse the lead in the solution.
4. Record the voltage measured by the voltmeter. If the voltage reading is negative, switch the leads with the metals to see a positive reading. Since all of the cell reactions in this experiment are spontaneous, all voltage readings should be positive.
5. Remove the leads from the metal strips.
6. Use the forceps to retrieve the salt bridge and dispose of it in the trash.

**4. Comparing the Zn/Zn2+ half-cell with the unknown concentration Cu/Cu2+ cell**

1. Use the forceps to retrieve one filter paper strip from the 1 M KCl solution. Allow excess solution to drip from the strip back into the container.
2. Use the forceps to drape the paper strip from the well containing the unknown M CuSO4 solution, across the top of the microwell plate, and down into the well containing the 1 M ZnSO­4 solution. This paper acts as the salt bridge between the two half-cells. Do not allow the paper to touch the metal strips.
3. Quickly touch the leads of the voltmeter to the copper strip in one well and to the zinc strip in the adjacent well. Try not to put too much pressure on the metal strips. This will cause them to fall and possibly touch the salt bridge. The best place to touch the strips is about half way down, being careful not to immerse the lead in the solution.
4. Record the voltage measured by the voltmeter. If the voltage reading is negative, switch the leads with the metals to see a positive reading. Since all of the cell reactions in this experiment are spontaneous, all voltage readings should be positive.
5. Remove the leads from the metal strips.
6. Use the forceps to retrieve the salt bridge and dispose of it in the trash.
7. Dispose of all solutions as directed by your instructor. Rinse and dry the metal strips, as they are reuseable.

**Calculations**

* 1. Use a Table of Standard Reduction Potentials to calculate the cell potential (as voltage) of the galvanic cell.
  2. Compare the Galvanic cell potential that you measured with the voltmeter to that calculated with the standard reduction potentials. Calculate the percent error.
  3. Use the voltages that you measured with the concentration cell (Procedure 3) and the Galvanic cell (Procedure 2) to calculate the molarity of Cu2+ in the unknown solution.
  4. Use the voltage that you measured with the cell in Procedure 4 and that of the Galvanic cell (Procedure 2) to calculate the molarity of Cu2+ in the unknown solution.

**Post-Lab Questions**

1. Write a balanced chemical equation to show the reaction that occurred in the Galvanic cell.
2. Write the line notation shorthand for the Galvanic cell.
3. Did the voltage that you measured with the voltmeter match the cell potential that you calculated from the standard reduction potential values? Explain why the two may not have matched.
4. For each of questions 4a – 4e, explain your reasoning. Predict the effect on the cell potential of the Galvanic cell if you
   1. increased the molarity of the Cu2+ ion
   2. increased the molarity of the Zn2+ ion
   3. poured K2CO3 solution into the CuSO4 solution
   4. increased the size of the metal strips
   5. increased the temperature to 35 °C
5. Did you find the same concentration for the unknown Cu2+ solution in Calculations 3 and 4? Would you expect to find the same value? Explain.
6. Based on your observations and calculations in this experiment, which is the more active metal, copper or zinc? Explain your reasoning.