**(Lab 5) Galvanic Cells:** electrons flowing spontaneously

Introduction

Most of our modern economy and civilization are based on moving electrons in time and space in a controlled manner. From the macroscale infrastructure of our electrical grid to the microscale of the circuitry in our mobile phones, we expect electrons to flow where we want them, when we want them.

We will investigate some fundamentals of electron flow in solution phase systems in this lab. Our emphasis will be on pairing a variety of half-cells to build circuits, and collecting data to characterize the behavior of our cells.

The term *galvanic cell* is used to describe a circuit involving the spontaneous transfer of electrons. In this context, “spontaneous” has the meaning that once we have mixed two or more chemical species, a reaction occurs without our further intervention; something in the reaction gains electrons and something else has lost electrons. Instead of mixing the oxidizing agent and reducing agent in the same solution, we will prepare a separate solution of each and connect them with a wire. In this way we will exert some measure of control over where the electrons flow.

The electrode that is supplying the electrons (and hence being oxidized) is called the *anode*, while the electrode receiving the electrons (and is itself reduced) is called the *cathode.* Other components in our circuit will include a multimeter to measure the cell *potential*, and a *salt bridge* to neutralize the ionic charge that would otherwise build up at each terminal.

Laboratory

*Hypothesis:*

Through careful observation we can characterize the behavior of galvanic cells and determine the concentration of an unknown solution sample.

*Experiments:*

Polish two strips of copper, one strip of zinc, one of magnesium, and one of iron using sandpaper or steel wool. Rinse these five metal strips with dilute (1.0 M) HNO3 into the proper waste container (**caution!**), and again with deionized water. These polished metals will be your electrodes.

Fill one small beaker ¾-full with 0.1 M Zn(NO3)2,a second ¾-full with 0.1 M CuSO4,a third ¾-full with 0.1 M MgSO4, and a fourth ¾-full with 0.1 M FeSO4. These solutions will be your half-cells. Place a copper electrode in the CuSO4 solution and the zinc electrode in the Zn(NO3)2 solution. Prepare a salt bridge by rolling a large piece of filter paper and saturating it with 0.1 M KNO3 solution. Place one end of the filter paper in each half cell. Connect two leads to the multimeter (one in the COMmon port, one in the V//mA port), then the probe end of each lead to the electrodes. If you obtain a negative potential, switch which electrode the leads are contacting. Set the multimeter to the appropriate range (~2000 mV is a good first guess). Record your measurement, and identify which electrode is the cathode and which is the anode. Repeat for the other five pairings of half-cells we can build using these four solutions and electrodes. Remember to prepare a fresh salt bridge for each new cell.

Construct a cell using your 0.1 M Zn(NO3)2 half-cell paired with the CuSO4 solution of unknown concentration, measure the cell potential and (assuming the reduction potential of the Zn2+(0.1 M)/Zn half-cell is 0.76 V) calculate the reduction potential of this unknown Cu half-cell relative to the Zn half-cell.

Starting with the 1.0 M CuSO4 solution provided, prepare a series of 10-fold dilutions to obtain CuSO4 solutions of 0.1, 0.01, 0.001, and 0.0001 M concentrations. Portions of ca. 35 mL of each solution should suffice for all experiments.

Set up a galvanic cell using 1.0 MCuSO4 and 0.001 M CuSO4, and immerse a polished copper electrode in each. Measure the cell potential and determine the cathode and anode. Write an equation for the reaction occurring in each half-cell.

Set up four more galvanic cells, this time using the Zn/Zn2+ half-cell each time, and pairing it with the 0.1 M. 0.01 M, 0.001 M, and 0.0001 M CuSO4 half-cells in sequence. Make sure to prepare a new salt bridge each time you change the concentration. You may need to reset the multimeter to a lower range for these experiments. Record the potential difference for each pairing.

*Data:*

-For the Cu/Zn, Mg/Zn, and Fe/Zn cells, assume the reduction potential of the Zn2+(0.1 M)/Zn half-cell is 0.76 V. Determine the reduction potential of the other three half-cells.

-For your Zn vs. Cu cells of varying [Cu2+], compare to the theoretical cell potential from a table of standard reduction potentials and the Nernst equation. How do things compare from your experimental values vs. theory? Your instructor will be interested to see your calibration curve plot.

*Conclusions:*

Can you write balanced equations for the oxidation/reduction reaction that took place in each cell you constructed? Can you identify the cathode and anode for each? What was the molar concentration of your unknown solution? For each cell you constructed, which direction did electrons flow? What about the direction of cations and anions flowing through the salt bridges? For the cell constructed using two copper electrodes, why did a current flow even though the electrodes were identical?