NYB

LAB EXERCISE 14

Galvanic Cell Measurements

Introduction:

Any oxidation-reduction reaction can be split up into two half-reactions, one showing the oxidation and the other the reduction. In the case of the redox reaction:

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

which occurs when a piece of metallic zinc is placed in a solution of copper(II) sulfate, the two half-reactions are:

$$Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2e^- \qquad \text{(oxidation)}$$
$$Cu^{2+}_{(aq)} + 2e^- \rightarrow Cu_{(s)} \qquad \text{(reduction)}$$

The tendency for this oxidation-reduction process to occur can be measured if the two reactions are allowed to occur in separate regions connected by a barrier that is porous to ion movement. Such an assembly is called a galvanic cell.

If a voltmeter is connected between the two electrodes, the voltmeter will show a voltage associated with the complete cell which is composed of the two half-cells. If the reduction potential of one of the half-cells is already known, then that of the other half-cell can be calculated. The usual convention is to arbitrarily assign a value of 0.00 V to the standard potential for the hydrogen electrode. On this scale, the reduction reaction below has a standard reduction potential of −0.763 V:

$$Zn^{2+}_{(aq)} + 2e^- \rightarrow Zn_{(s)}$$

The purpose of this experiment is to measure the voltage of three galvanic cells: $Zn|Zn^{2+}$ with $Cu|Cu^{2+}$; $Zn|Zn^{2+}$ with $Pb|Pb^{2+}$; and $Cu|Cu^{2+}$ with $Pb|Pb^{2+}$. Since the cell voltage will not be measured at standard concentrations (1 mol/L for both ions), the standard cell potential, E°, must be obtained by using the Nernst equation:

$$E = E° - \frac{0.0592}{n} \log Q$$

Knowing this, and taking the standard reduction potential of $Zn|Zn^{2+}$ as −0.763 V, the standard reduction potentials of the following half-reactions will be calculated:

$$Cu^{2+}_{(aq)} + 2e^- \rightarrow Cu_{(s)}$$
$$Pb^{2+}_{(aq)} + 2e^- \rightarrow Pb_{(s)}$$

The effect of concentration changes on the overall cell voltage will also be measured. These experimental results can be compared with the results expected from the Nernst equation.

Pre-Lab Questions

What is "Q" in the Nernst equation?
Under what circumstances/conditions does the Nernst equation need to be used?
How can the Nernst equation be used to calculate the concentration of the unknown solution?

Procedure

Rinse the electrodes and the cells with distilled water and dry them with a piece of paper towel. Pour 30 mL of 0.10 mol/L zinc sulfate ($ZnSO_4$) solution into the porous cup. Pour 50 mL of 0.10 mol/L copper(II) sulfate ($CuSO_4$) solution into the outer glass container. Carefully dip the zinc electrode into the Zn^{2+} solution and the copper electrode into the Cu^{2+} solution. Connect the two electrodes to the voltmeter and measure the cell voltage. *If by mistake the polarity is wrong, then the voltage will be negative. If this happens, immediately turn the Operation switch to "standby", and exchange the connections at the two electrodes*.

After the reading has been taken turn the Operation switch to "standby". Wait one minute and then measure the voltage again. Take the average of the two readings. As soon as the voltage measurement is finished, turn the Operation switch to "standby" and prepare for the next measurement.

Transfer the copper(II) sulfate solution to a clean, dry beaker for later use. Remove the copper electrode. Rinse and dry the glass container and the electrode. Pour 50 mL of the lead(II) sulfate ($PbSO_4$) solution into the glass container. Set up the lead half-cell. Put back the zinc half-cell and measure the cell voltage.

Transfer the zinc sulfate solution to a clean, dry beaker. Rinse and dry the porous cup and the zinc electrode. Put 30 mL of the 0.10 mol/L copper(II) sulfate solution stored in the beaker from the previous experiment into the porous cup. Set up the copper half-cell, combine it with the half-cell and measure the cell potential. Dismantle, clean and dry the apparatus.

Set up the following cell and measure its potential:

$$Zn \mid Zn^{2+}\ (0.10\ mol/L) \mid Cu^{2+}\ (0.0010\ mol/L) \mid Cu$$

Make up the Cu^{2+} solution by dilution of the stock solution supplied.

Set up the following cell using an assigned Cu^{2+} solution of unknown concentration and measure its potential:

$$Zn \mid Zn^{2+}\ (0.10\ mol/L) \mid Cu^{2+}\ (?\ mol/L) \mid Cu$$

Calculate the unknown Cu^{2+} concentration.