EXPERIMENT 10
Electrochemical Cells

Introduction
An electrochemical cell is a device that may be used for converting chemical energy into electrical energy. An oxidation-reduction reaction is the basis for designing an electrochemical cell. The tendencies of metals to be oxidized differently are often used to design an electrochemical cell based on their reaction differences. Oxidation-reduction or redox reactions involve the transfer of electrons from one reactant to another. Oxidation and reduction half-reactions each represent half of the overall reaction. When adding the half-equations to yield the overall equation, the electrons appearing in the two half-cell reactions must cancel each other. In any oxidation-reduction reaction equation, the number of electrons released must be equal to the number of electrons consumed. For example, the oxidation of zinc metal and reduction of Cu$^{2+}$ for a spontaneous electrochemical reaction results in the transfer of 2 electrons.

\[
\begin{align*}
\text{Zn} (s) & \rightarrow \text{Zn}^{2+} (aq) + 2e^- & \text{oxidation half-reaction} \\
\text{Cu}^{2+} (aq) + 2e^- & \rightarrow \text{Cu} (s) & \text{reduction half-reaction} \\
\text{Zn} (s) + \text{Cu}^{2+} (aq) & \rightarrow \text{Zn}^{2+} (aq) + \text{Cu} (s) & \text{overall reaction}
\end{align*}
\]

This reaction can be observed by placing a strip of zinc metal into a solution of copper(II) sulfate. As the reaction goes forward, metallic copper can be seen being deposited on the zinc surface. Eventually, the blue color characteristic of Cu$^{2+}$ ions in solution will fade and metallic copper will begin to precipitate out of the solution. The strip of zinc metal is gradually being consumed during this process, indicating zinc atoms are being oxidized to zinc (II) ions. This observation shows that zinc is more easily oxidized than copper. Alternatively, the Cu$^{2+}$ ion can be described as being more likely to accept electrons from Zn metal than Zn$^{2+}$ ion is to accept electrons from Cu metal; a Cu$^{2+}$ ion is easier to reduce than Zn$^{2+}$ and thus has a greater reduction potential. The standard reduction potentials, E° red, for metal ions provide a quantitative measure of a metal ion’s tendency to accept electrons as seen in Table 1.

<table>
<thead>
<tr>
<th>Reduction Half Equation</th>
<th>E° red</th>
<th>V</th>
</tr>
</thead>
<tbody>
<tr>
<td>$\text{Al}^{3+} (aq) + 3 , e^- \rightarrow \text{Al} (s)$</td>
<td>-1.66</td>
<td></td>
</tr>
<tr>
<td>$\text{Zn}^{2+} (aq) + 2 , e^- \rightarrow \text{Zn} (s)$</td>
<td>-0.76</td>
<td></td>
</tr>
<tr>
<td>$\text{Fe}^{2+} (aq) + 2 , e^- \rightarrow \text{Fe} (s)$</td>
<td>-0.44</td>
<td></td>
</tr>
<tr>
<td>$\text{Sn}^{2+} (aq) + 2 , e^- \rightarrow \text{Sn} (s)$</td>
<td>-0.14</td>
<td></td>
</tr>
<tr>
<td>$\text{Pb}^{2+} (aq) + 2 , e^- \rightarrow \text{Pb} (s)$</td>
<td>-0.13</td>
<td></td>
</tr>
<tr>
<td>$2 , \text{H}^+ (aq) + 2 , e^- \rightarrow \text{H}_2 (g)$</td>
<td>0.00</td>
<td></td>
</tr>
<tr>
<td>$\text{Cu}^{2+} (aq) + 2 , e^- \rightarrow \text{Cu} (s)$</td>
<td>+0.34</td>
<td></td>
</tr>
<tr>
<td>$\text{Ag}^+ (aq) + 2 , e^- \rightarrow \text{Ag} (s)$</td>
<td>+0.80</td>
<td></td>
</tr>
</tbody>
</table>

Using Standard Reduction Potential Tables
The potentials in the reduction potential tables are determined under standard conditions, usually 1M concentrations, 25°C, and 1 atm. The reference point for these potentials is the reduction potential of hydrogen ions (H$^+$) in aqueous solution being converted to hydrogen gas (H$_2$). There are two main ideas to remember when using these tables.

1. The more positive the reduction potential, the easier it is to reduce the ion to its metallic form.

2. The oxidation half-reactions of the metals in the table are simply the reverse of the half-reactions. The standard oxidation potential, E° oxid of an oxidation half-reaction has the same numerical value, but the opposite sign of the corresponding standard reduction potential.

The reduction potential table is used to make predictions about the spontaneity for oxidation-reduction reactions involving metals in the table. For example, if you place a Zn strip in a solution of CuSO$_4$ as done above, you would be able to predict whether the reaction was spontaneous by calculating the overall cell voltage for the reaction. First you must determine both the oxidation and reduction reactions. Since you know that the zinc is in the metallic form in the strip, you are predicting it will be oxidized to its ionic form, Zn$^{2+}$(aq).

\[
\begin{align*}
\text{Zn} (s) & \rightarrow \text{Zn}^{2+} (aq) + 2e^- & \text{oxidation half-reaction} \\
E°_{\text{oxid}} & = E°_{\text{red}} = +0.76V
\end{align*}
\]
The copper ion in solution must then gain electrons as it is reduced to copper metal.

\[
\text{Cu}^{2+} (aq) + 2e^- \rightarrow \text{Cu} (s) \quad E^\circ_{\text{red}} = +0.34 \text{ V}
\]

By adding the reactions together, you now have both an oxidation and a reduction. Simply add the voltages of these two reactions to determine the overall cell voltage. If the value is positive, the reaction is spontaneous; you will see the zinc strip dissolve and solid copper precipitate. If the value is negative, nothing will happen, indicating a non-spontaneous reaction.

\[
E^\circ_{\text{oxid}} + E^\circ_{\text{red}} = E^\circ_{\text{cell}}
\]

\[
\text{Zn} (s) + \text{Cu}^{2+} (aq) \rightarrow \text{Zn}^{2+} (aq) + \text{Cu} (s) \quad E^\circ_{\text{cell}} = 1.10 \text{ V}
\]

The positive cell voltage calculated from the standard reduction potential tables indicates that the reaction is spontaneous, which is why you will observe the dissolution of the zinc and the production of copper metal when the zinc strip is immersed in the copper (II) sulfate solution. In writing a balanced net cell reaction for metals that may not exchange the same number of electrons, we may need to multiply the coefficients of the half-cell reactions to properly cancel the electrons. However, changing the coefficients of a half-cell reaction does not influence the reduction and oxidation potential. Reduction and oxidation potentials are intensive properties in that they are independent of the amount of substance, so there is no need to adjust the values of the voltages for the number of electrons in the reaction. The voltage is only dependent on the metals involved, not the stoichiometry of the chemical equation.

A voltaic cell, also known as a Galvanic cell, is an electrochemical cell where a spontaneous reaction generates an electrical current. A picture of a typical electrochemical cell is shown in Figure 1. A voltaic cell consists of two connected half-cells, one containing the anode and the other the cathode. The connection allows a path for electrons to flow from one metal electrode to another through an external circuit and an internal cell connection (or salt bridge). Oxidation, or loss of electrons, occurs at the anode, while reduction, or gain of electrons, occurs at the cathode. In this experiment you will build various electrochemical cells using Cu, Zn, Pb, and Fe and measure the voltages. Based upon your observed voltages, you will be able to rank the metals in order of their relative ease of oxidation and compare the measure cell voltages with those calculated from the standard reduction potentials.

Figure 1:
Experimental Procedure
Your TA must sign your data before leaving lab or you will not receive credit for attending the lab.
Note: Metal strips should appear bright and shiny before beginning each section of the lab.
Look carefully for changes either in the appearance of the metal surfaces or the solutions to indicate a reaction.

Part 1: The Zn-Cu Redox Reaction
1. Clean the Zn metal with sandpaper.
2. Place the clean metal in a small test tube.
3. Fill the test tube with 0.1M CuSO$_4$ solution; be sure that the Zn metal is completely submerged.
4. After 2.5 minutes, record your observations.
5. After 5 minutes, record your observations.
6. Clean the Zn metal strip with sandpaper and dry using paper towels and return to the container.

Part 2: The Pb-Cu Redox Reaction
1. Clean the Pb metal with sandpaper.
2. Place the clean metal in a small test tube.
3. Fill the test tube with 0.1M CuSO$_4$ solution; be sure that the Pb metal is completely submerged.
4. After 2.5 minutes, record your observations.
5. After 5 minutes, record your observations.
6. Clean the Pb metal with sandpaper and dry using paper towels and return to the container.
7. Pour the CuSO$_4$ solution into waste beaker.

Part 3: The Zn-Pb Redox Reaction
1. Clean the Zn metal with sandpaper.
2. Place the clean metal in a small test tube.
3. Fill the test tube with 0.1M Pb(NO$_3$)$_2$ solution; be sure that the Zn metal is completely submerged.
4. After 2.5 minutes, record your observations.
5. After 5 minutes, record your observations.
6. Clean the Zn metal strip with sandpaper and dry using paper towels and return to the container.
7. Discard the Pb(NO$_3$)$_2$ solution into waste beaker.

Part 4: Electrochemical Half-Cell Reactions
1. Dispense 10 mL of solutions 0.1 M CuSO$_4$, 0.1 M Zn(NO$_3$)$_2$, 0.1 M Pb(NO$_3$)$_2$, 0.1 M FeSO$_4$ and 0.1 M KNO$_3$ into 30 mL labeled beakers.
2. Clean the copper, zinc, lead, and iron electrodes using sandpaper and rinse with deionized water.
3. Place each metal electrode in its corresponding ionic solution; e.g. copper strip goes into the CuSO$_4$ solution. It is important that the correct metal is in the correct solution or your cell will not work properly.
4. Obtain small strips of filter paper to be used as salt bridges. Completely wet one strip in the beaker containing 0.1 M KNO$_3$.
5. Carefully remove the completely wet strip and place one end in the CuSO$_4$ solution and the other in the Zn(NO$_3$)$_2$ solution. The salt bridge should not touch the electrodes.
6. Attach one alligator clip from the positive terminal on the voltmeter to the Cu electrode and the second clip on the negative terminal to the Zn electrode.
7. Record the voltage of the electrochemical cell. Record the copper as the cathode and the zinc as the anode.
8. Repeat for the remaining cells in the data table, recording the positive cell voltages and the anode (- terminal) and cathode (+ terminal) for each cell. Use a new wet piece of filter paper as a salt bridge for each determination.
9. Clean the metal strips with sandpaper and dry each strip using paper towels.
10. Dispose of all solutions into the appropriate waste container.
11. Return all metal pieces to their original container clean and dry.
12. Clean and dry your work area with water. Wash all glassware with soap then rinse 3 times with tap water, and once with deionized water.
Table 1: Recorded Electrochemical Potentials

<table>
<thead>
<tr>
<th>Cell</th>
<th>Measured $E_{\text{cell}}$</th>
<th>Anode</th>
<th>Cathode</th>
</tr>
</thead>
<tbody>
<tr>
<td>Cu-Zn</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Cu-Pb</td>
<td></td>
<td></td>
<td></td>
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<tr>
<td>Cu-Fe</td>
<td></td>
<td></td>
<td></td>
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<tr>
<td>Zn-Pb</td>
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<td></td>
<td></td>
</tr>
<tr>
<td>Zn-Fe</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Pb-Fe</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
Calculations

**Overall Cell Reactions for Parts 1-3**
First, identify which metal is reduced and which is oxidized based on your experimental parameters. For example, in part 1, zinc metal and copper ion were the reactants. These will also need to be the reactants in your overall cell reaction, so for an oxidation-reduction reaction to occur, you must have the Zn(s) being oxidized to Zn^{2+}(aq) and the Cu^{2+}(aq) being reduced to Cu(s). Write the oxidation reaction, the reduction reaction and the overall cell reaction for each of the metal combinations for parts 1 through 3.

**Overall Cell Reactions for Part 4**
First, identify which metal is reduced and which is oxidized based on your identification of the cathode and anode on your data sheet. Reduction takes place at the cathode and oxidation at the anode. Write the overall cell reaction for each of the metal combinations for part 4.

**Calculating Cell Potentials for Part 4**
Locate both metals in the reduction potential table. Identify which of the two metals is oxidized. Reverse the sign of the reduction reaction for that metal to determine the oxidation potential. Then, calculate the cell voltage, $E^{\circ}_{\text{cell}}$, by adding the value of the voltage for the reduction to the value for the oxidation.

$$E^{\circ}_{\text{oxid}} = -E^{\circ}_{\text{red}}$$

$$E^{\circ}_{\text{cell}} = E^{\circ}_{\text{oxid}} + E^{\circ}_{\text{red}}$$

**Percent Difference between the Calculated Cell Potential and the Measured Cell Potential**
Use the equation below to calculate the percent difference between the calculated and measured cell potentials.

$$\% \text{Difference} = \left| \frac{E_{\text{cell(measured)}} - E_{\text{cell(calculated)}}}{E_{\text{cell(calculated)}}} \right| \times 100$$
Results

Hand in the following results
Include the correct number of significant figures for each measurement.

<table>
<thead>
<tr>
<th>Cell</th>
<th>Half Reaction (Oxidation)</th>
<th>Half Reaction (Reduction)</th>
<th>Overall Cell Reaction</th>
</tr>
</thead>
<tbody>
<tr>
<td>Cu-Zn</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Cu-Pb</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Zn-Pb</td>
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<td></td>
<td></td>
</tr>
</tbody>
</table>

Table 3: Cell Reactions for Part 4

<table>
<thead>
<tr>
<th>Cell</th>
<th>Overall Cell Reaction</th>
<th>Measured $E_{\text{cell}}$</th>
<th>Calculated $E_{\text{cell}}$</th>
<th>% Deviation</th>
</tr>
</thead>
<tbody>
<tr>
<td>Cu-Zn</td>
<td></td>
<td></td>
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<tr>
<td>Cu-Pb</td>
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<td>Cu-Fe</td>
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<td>Zn-Pb</td>
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<tr>
<td>Pb-Fe</td>
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</table>
Discussion Questions

You must use your data to answer these questions. If the answer you give does not reflect YOUR data and results, you will receive no credit for the question.

1. Were your cell reactions in part 1 spontaneous or non-spontaneous? Briefly explain for each reaction.

2. Calculate the cell potentials for each of the reactions in parts 1-3. Do the calculations support the conclusions you made about the spontaneity made in question 1? Briefly explain.

3. Were your percent differences between the measured $E^\circ_{\text{cell}}$ and the calculated $E^\circ_{\text{cell}}$ less than 5% for all of your reactions? If not, give at least 2 specific reasons that the experimental results would vary from the tabulated $E^\circ_{\text{cell}}$.

4. Use the table of standard cell potentials to predict whether an electrochemical cell containing tin ion, Sn$^{2+}$ with a tin electrode, and a lead electrode in a solution of lead nitrate, Pb(NO$_3$)$_2$, would produce solid tin or lead. Show calculations to support your conclusion.

5. Write the overall cell reaction for a cell containing aluminum metal in a solution of zinc(II)sulfate. Calculate the cell potential for this reaction. Is the reaction spontaneous or non-spontaneous?
The penalty for not paying your lab bill is an incomplete. If you do have a bill, you will receive an incomplete and be dropped from any registered labs in the chemistry department until the bill is paid. You can check your bill status with the stockroom (room 210).

Makeup labs are held April 26th and 27th. You need to sign up in the stockroom before being allowed to attend a makeup lab. Your final will be held in your lab the week of April 18-22. You do not need any safety equipment to take the final.

1. What is the penalty if you don’t pay your lab bill?

2. Where do you sign up for a makeup lab?

3. When and where is your final in this course? Give the date, time and location.

Answer the following questions based on the following electrochemical cell:

\[
\text{Zn}(s) / \text{Zn}^{2+}(aq) \rightleftharpoons \text{Sn}^{2+}(aq) / \text{Sn}(s) \quad \text{Zn}^{2+}(aq) + 2 \, \text{e}^- = \text{Zn}(s) \quad E = -0.76V
\]

\[
\text{Sn}^{2+}(aq) + 2\, \text{e}^- = \text{Sn}(s) \quad E = -0.14V
\]

4. Write the chemical equation associated with this cell.

5. How many electrons are being transferred in the reaction?

6. Calculate the cell potential.

7. Is the reaction spontaneous or nonspontaneous?

8. Which species has been oxidized?

9. Identify the anode in the cell.

10. Which electrode is being dissolved?