EXPERIMENT 5

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.

$Zn\ (s) \rightarrow Zn^{2+}\ (aq) + 2e^-$	**oxidation half-reaction**
$Cu^{2+}\ (aq) + 2e^- \rightarrow Cu\ (s)$	**reduction half-reaction**
$Zn\ (s) + Cu^{2+}\ (aq) \rightarrow Zn^{2+}\ (aq) + Cu\ (s)$	***overall reaction***

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 **voltage** that occurs when a reaction occurs can be measured and is called the **cell potential.** Using a series of cell potentials, a table of **standard reduction potentials, $E°_{red}$**, for metal ions was developed. This table provides a quantitative measure of a metal ion's tendency to accept electrons.

Table 1: Standard Reduction Potentials

Reduction Half Equation $E°_{red}$	**V**
Al^{3+} (aq) + 3 e^- → Al (s)	-1.66
Zn^{2+} (aq) + 2 e^- → Zn (s)	-0.76
Fe^{2+} (aq) + 2 e^- → Fe (s)	-0.44
Sn^{2+} (aq) + 2 e^- → Sn (s)	-0.14
Pb^{2+} (aq) + 2 e^- → Pb (s)	-0.13
2 H^+ (aq) + 2 e^- → H_2 (g)	**0.00**
Cu^{2+} (aq) + 2 e^- → Cu (s)	+0.34
Ag^+ (aq) + 1 e^- → Ag (s)	+0.80

Using Standard Reduction Potential Tables

The standard potentials in the reduction potential tables are the voltages determined for the chemical reaction 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 relative **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). But is this true or does the zinc want to remain in its metallic form.

Here is the overall chemical reaction describing the experiment:

$$\mathbf{Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)}$$

The copper ion in solution must gain electrons as it is reduced to copper metal.

$\mathbf{Cu^{2+}\ (aq) + 2e^- \rightarrow Cu\ (s)}$ **reduction half-reaction**

Reduction table data is for $Cu^{2+}\ (aq) + 2e^- \rightarrow Cu\ (s)$ will be given as $E°_{Cu} = +0.34$ V

The solid zinc must lose electrons and dissolve to form Zn^{2+} ions, so it is oxidized.

$\mathbf{Zn\ (s) \rightarrow Zn^{2+}\ (aq) + 2e^-}$ **oxidation half-reaction**

Reduction table data is for $Zn^{2+}\ (aq) + 2e^- \rightarrow Zn\ (s)$ will be given as $E°_{Zn} = -0.76V$

By adding the reactions together, you now have both an **oxidation** and a **reduction**. Use the equation below to determine the 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.

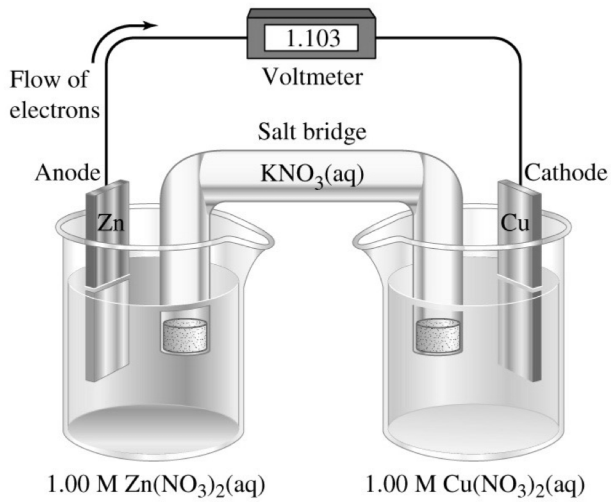
$\mathbf{E°_{cell} = E°_{red} - E°_{oxid}}$ so you would use the following values: $\mathbf{E°_{cell} = E°_{Cu} - E°_{Zn}}$

$$\mathbf{Zn\ (s) + Cu^{2+}\ (aq) \rightarrow Zn^{2+}\ (aq) + Cu\ (s)\quad E°_{cell} = +0.34V - (-0.76V) = +1.10\ V}$$

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**). The measurement of the voltage is done using a voltmeter inserted within the external circuit. **Oxidation**, or loss of electrons, occurs at the **anode**, while **reduction**, or gain of electrons, occurs at the **cathode**. If this were a nonspontaneous cell, (**electrolytic cell**), the voltage measured would be negative, indicating that the flow of electrons is occurring in the opposite direction.

Figure 1:



Balancing electrochemical reactions

In the previous examples, each reactant was transferring 2 electrons between the atoms. This is not always the case. If you have differing charges on the ions, you will have to take that into account when the reaction is balanced. In the example below, by balancing with 3 zinc atoms to 2 aluminum atoms, the overall charge on each side is 6+. Using half-reactions can help you see how the electrons will balance.

$\mathbf{Cu^{2+}}$ ***(aq)*** **+ 2e⁻ → Cu** ***(s)*** **reduction half-reaction:** $\mathbf{E°_{Cu} = +0.34\ V}$

The solid aluminum must lose electrons and dissolve to form Al^{3+} ions, so it is oxidized.

$\mathbf{Al\ (s) \rightarrow Al^{3+}\ (aq) + 3e^-}$ **oxidation half-reaction** $\mathbf{E°_{Al} = -1.66\ V}$

$$E°_{cell} = +0.34V - (-1.66V) = +2.00\ V$$

For the balanced reaction, the number of electrons must match on both sides, so cross multiply each half-reaction with the number of electrons from the other half-reaction. When the 2 half-reactions are added together, the electrons will then cancel out on each side of the reaction.

$$3(Cu^{2+}\ (aq) + 2e^- \rightarrow Cu\ (s)) \quad = \quad 3Cu^{2+}\ (aq) + 6e^- \rightarrow 3Cu\ (s))$$

$$+\ \underline{2(Al\ (s) \rightarrow Al^{3+}\ (aq) + 3e^-) \quad = \quad 2Al\ (s) \rightarrow 2Al^{3+}\ (aq) + 6e^-}$$

$$\cancel{6e-} + 2Al(s) + 3Cu^{2+}(aq) \rightarrow 2Al^{3+}(aq) + 3Cu(s) + \cancel{6e-}$$

In Your Lab…

In this experiment you will build various electrochemical cells using Cu, Zn, Al, 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.

Experiment 5: Prelab Worksheet
(Submit through Brightspace before the start of your lab session)

Name: ________________________________ Date: ________ Section: ________ Grade: ________

Record all values with the correct number of significant figures and units.
Place all answers on the line next to the question.
Show calculations for any numerical answers.
See any 114 TA in the help office before your prelab is due if you have any questions.
Your answer must be completely correct to get any credit for the answer, no partial credit.

Consider an electrochemical cell constructed from the following half cells, linked by an external circuit and by a KCl salt bridge.

- an Al(s) electrode in 1.0 M $Al(NO_3)_3$ solution
- a Pb(s) electrode in 1.0 M $Pb(NO_3)_2$ solution

1. Write the **balanced** overall cell reaction that will make the cell spontaneous. Keep in mind that aluminum will transfer 3 electrons per aluminum atom while lead will only transfer 2.

2. Calculate $E°_{cell}$

3. How many electrons will be transferred in the balanced overall reaction?

4. Sketch a quick outline of the cell based on the figure given in the introduction.

For the following reaction: $Al(s) + 3Ag^{+}(aq) \rightarrow Al^{3+}(aq) + 3Ag(s)$

5. Calculate $E°_{cell}$. Include the sign with your answer.

6. Is the reaction spontaneous or nonspontaneous?

7. Identify the cathode.

8. Identify the reducing agent.

9. Write the half-reaction for the oxidation.

10. How many electrons will be transferred in the OVERALL reaction?

Experiment 5: Experimental Procedures and Data Sheet
Submit as part of your informal report

Name: ____ Date: ________ Section: ________

TA Signature: ______________________________________

All data must be written in pen at the time it is collected. **Pencil is not allowed!!**
Record all measurements with the correct number of significant figures and units.
TA signature and TA initials on any changes made to the data must be present or your data is invalid

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 if it is not shiny.
2. Place the clean metal in a small beaker.
3. Fill the beaker with enough 0.1M $CuSO_4$ solution until the Zn metal is partially submerged.
4. After 2.5 minutes, record your observations.

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5. After 5 minutes, record your observations.

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6. Clean the Zn metal strip with sandpaper and dry using paper towels and return to the container.
7. Retain the $CuSO_4$ solution for use in Part 2.

Part 2: The Al-Cu Redox Reaction

1. Clean the aluminum metal if not shiny.
2. Place the clean metal in a small beaker.
3. Fill the beaker with enough 0.1M $CuSO_4$ solution that the Al metal is partially submerged.
4. After 2.5 minutes, record your observations.

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5. After 5 minutes, record your observations.

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6. Clean the Al 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-Al Redox Reaction

1. Clean the Zn metal with sandpaper.
2. Place the clean metal in a small beaker.
3. Fill the beaker with enough 0.1M $Al(NO_3)_2$ solution that the Zn metal is partially submerged.
4. After 2.5 minutes, record your observations.

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5. After 5 minutes, record your observations.

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6. Clean the Zn metal with sandpaper and dry using paper towels and return to the container
7. Discard the $Al(NO_3)_2$ 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 $Al(NO_3)_2$, 0.1 M $FeSO_4$ and 0.1 M KNO_3 into 30 mL labeled beakers.
2. Clean the copper, zinc 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 (red) on the voltmeter to the Cu electrode and the second clip on the negative terminal (black) 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 cell voltages for each cell. Be sure to note if the voltage is positive or negative. Use a new wet piece of filter paper as a salt bridge for each determination. Record the metal attached to the red terminal as the cathode and the metal attached to the black terminal is the anode.
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.

Cell Reactions for Part 4

Cell	**Anode**	**Cathode**	**Cell Reaction**	**Measured E_{cell}**
Cu-Zn				
Cu-Al				
Cu-Fe				
Zn-Al				
Zn-Fe				
Al-Fe				

Grading

Points		
Neatness and Clarity of Data	5pts	___________pts
Significant figures and units	5pts	___________pts
All data is present	10pts	___________pts
Deductions (sliding based on TA discretion)		
Lab area left unclean	20pts	___________pts
Improper waste disposal	20pts	___________pts
Disruptive behavior	20pts	___________pts
Other: ____		___________pts
Plagiarism!!! Data are identical to another student	**100pts**	**__________pts**

Grade for Experimental Procedures and Data ________pts

Experiment 5: Results Table

Submit as part of your informal report

Name: ____ Date: ________ Section: ________

All data must be written in pen at the time it is collected. **Pencil is not allowed!!**
Record all results with the correct number of significant figures and units
No stray marks or notes should be present on this page. Only the tabulated results are allowed

Table 2: Cell Reactions for Parts 1-3

Cell	Half Reaction (Reduction)	Half Reaction (Oxidation)	Cell Reaction
Cu-Zn			
Cu-Al			
Zn-Al			

Table 3: Cell Reactions for Part 4

Cell	Metal Reduced	Metal Oxidized	Cell Reaction	Calculated E_{cell}	Measured E_{cell}	% Difference
Cu-Zn						
Cu-Al						
Cu-Fe						
Zn-Al						
Zn-Fe						
Al-Fe						

Grading

Points

Significant figures and units	5pts	__________pts
Table is neat and legible	5pts	__________pts
All results are present	10pts	__________pts

Deductions (sliding based on TA discretion)

Results do not make sense	20pts	__________pts
Results do not match data	20pts	__________pts
Other: ____		__________pts

Results table __________pts

Calculations Section: Submit as part of your informal report

You must be able to perform these calculations with given data on your concept review quiz.
Fill in the results table as you go along for any results that do not require 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, $\mathbf{E°_{cell}}$, by adding the value of the voltage for the reduction to the value for the oxidation.

$$\mathbf{E°_{cell} = E°_{red} - E°_{oxid}}$$

Cell	Calculated $E°_{cell}$
Cu-Zn	
Cu-Al	
Cu-Fe	
Zn-Al	
Zn-Fe	
Al-Fe	

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.

$$\%Difference = \frac{\left| E_{cell(measured)} - E_{cell(calculated)} \right|}{E_{cell(calculated)}} x100$$

Cell	% Difference
Cu-Zn	
Cu-Al	
Cu-Fe	
Zn-Al	
Zn-Fe	
Al-Fe	

Experiment 5 Additional Questions:
Submit as part of your informal report

You should be able to do the following:

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

a.) Part 1 Zn-Cu

b.) Part 2 Al-Cu

c.) Part 3 Zn-Al

d.) Part 4 Cu-Zn

e.) Part 4 Cu-Al

f.) Part 4 Cu-Fe

g.) Part 4 Zn-Al

h.) Part 4 Zn-Fe

i.) Part 4 Al-Fe

2. Were your percent differences between the measured E°_{cell} and the calculated E°_{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°_{cell}.