Voltaic Cells

Objectives

- Determine the standard cell potentials for multiple cells
- Determine cell potential for non-standard cells
- Compare experimentally determined cell potentials to calculated potentials

Background

Electrochemistry

When electrons transfer between reaction components in a redox reaction, we can harness the motion of the electrons to create a potential. Electrochemistry revolves around the separation of the two half-reactions in a redox reaction and establishing two different electrodes. This might involve physically separating the half-reactions or including a separator, such as a semi-permeable membrane or plastic dividers.

With the reactions separated, the electrons will need to flow through the wire connecting the electrodes, creating a current. The spontaneous flow of electrons in this set-up is known as a voltaic or galvanic cell. This is the basic idea behind a battery. The cell generates electrical energy as the electrons spontaneously move from the negative electrode to the positive electrode.

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A cell with a spontaneous flow of electrons is known as a voltaic or galvanic cell, with the set-up as shown in Figure 1. The negative electrode is where the oxidation occurs and is known as the anode. Because oxidation is the loss of electrons, the anode is where electrons flow from. In the voltaic cell shown, the beaker with the zinc metal and zinc solution is the anode. Electrons flow toward the cathode where reduction occurs. In this cell, the beaker with the copper ions and copper metal is the cathode.

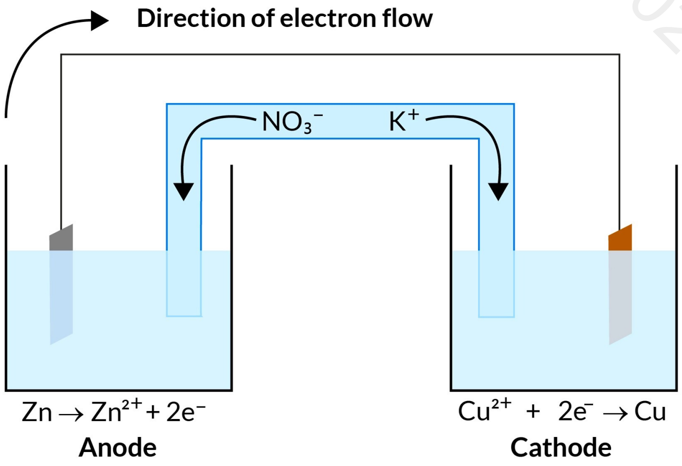

Figure 1 Schematic diagram of a voltaic cell with a zinc anode and a copper cathode

To complete a cell, there needs to be an exchange of ions to counteract the changing charge of electron transfer. One approach to this is a salt bridge, which allows spectator ions to move between the beakers. In a voltaic cell, anions will move toward the anode to counteract the growing number of cations. Cations will move through the salt bridge toward the cathode to counteract the decreasing number of cations as they are reduced.

$$\mathrm{Zn|\ Zn^{2+}(1.0\ M)||Cu^{2+}(1.0\ M)|Cu} \qquad \text{Equation 1}$$

The notation to represent a cell, shown in Equation 1, follows the flow of electrons from left to right so the anode appears first with the reactant and product of the half-reaction with a vertical line between them. Then there is a set of double vertical lines that indicates the separation from the next electrode. Finally, there appears the reactant and product of the reduction half-reaction at the cathode. The concentration of each solution involved is also noted.

Reduction Potentials

A voltaic cell creates a potential of usable electrical energy. Potential is measured in volts. One volt corresponds to 1 Joule of energy per Coulomb, which is a unit of charge. A spontaneous reaction that generates energy has a positive potential. To determine the potential of a cell, we need to look at the potential of each half-reaction involved, given as reduction potentials in Table 1.

Table 1 Standard reduction potentials used in the experiment

Reduction half-reaction	E^o_{red} (V)
$\mathrm{Ag^+(aq)} + e^- \rightarrow \mathrm{Ag(s)}$	+0.799
$\mathrm{Cu^{2+}(aq)} + 2e^- \rightarrow \mathrm{Cu(s)}$	+0.337
$\mathrm{Pb^{2+}(aq)} + 2e^- \rightarrow \mathrm{Pb(s)}$	–0.126
$\mathrm{Zn^{2+}(aq)} + 2e^- \rightarrow \mathrm{Zn(s)}$	–0.763

Half-reaction potentials are only given for the reactions written as reductions. The standard reduction potentials are usually measured against the standard hydrogen electrode. The half-reactions with positive reduction potential will occur as written. The half-reactions with negative reduction potentials are more likely to occur as oxidations.

$$E^o_{cell} = E^o_{red}(\text{cathode}) - E^o_{red}(\text{anode}) \qquad \text{Equation 2}$$

The standard potential of a cell is the potential difference between the reduction potentials for the two half-reactions, as shown in Equation 2. Subtract the reduction potential at the anode where electrons start from the reduction potential at the cathode where electrons end up.

Cell Conditions

The conditions of a cell determine its potential. A standard potential is one where the temperature is 298 Kelvin and all the involved solutions have a concentration of 1 molar. The standard condition is denoted by a degree symbol. Although potentials are reported this way, we can calculate systems under non-standard conditions of temperature and concentration using the Nernst equation, Equation 3.

$$E_{cell} = E^{o}_{cell} - \frac{RT}{nF} lnQ \qquad \text{Equation 3}$$

The Nernst equation is a correction factor to the standard cell potential. The correction factor involves two constants, the gas constant, R, and Faraday's constant, F. Temperature appears as T and must be included in units of Kelvin. Then, there are two components related to the balanced redox reaction. The n in the equation refers to the number of electrons transferred in the reaction, which is the number of electrons that appears in each half-reaction when the overall reaction is balanced. Q is the reaction quotient, which is the same expression as the equilibrium expression and only includes solutions or gases, not solids or liquids. The reaction quotient changes based on the concentrations of solutions present.

Materials

- Zinc strip
- Copper strip
- Lead strip
- Silver strip
- 1.0 M $Zn(NO_3)_2$
- 0.1 M $Zn(NO_3)_2$
- 1.0 M $Cu(NO_3)_2$
- 0.1 M $Cu(NO_3)_2$
- 1.0 M $Pb(NO_3)_2$
- 0.1 M $Pb(NO_3)_2$
- 0.1 M $AgNO_3$
- 1.0 M KNO_3
- Microwells
- Strips of filter paper

Safety goggles are required!

Use the multimeter and probes only as directed.

Procedure

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1. Connect the red (cathode) and black (anode) probes to the multimeter to read the DC voltage.
2. Set up the well-plate module. If using a circular well plate, fill the center well with 1.0 M KNO_3 and placing each of the remaining solutions in the designated outer wells, making sure to fill the inner channel. For other well-plates, prepare filter strips soaked in KNO_3 solution to act as salt bridges. **Make sure you know which well contains each solution.**
3. Take a piece of clean zinc foil and clip it to the black probe of the multimeter. Black will always be your anode.
4. Place the zinc foil into the 1.0 M $Zn(NO_3)_2$ solution.
5. Clip a piece of clean copper wire to the red probe of the multimeter.
6. Place the copper wire in the 1.0 M $Cu(NO_3)_2$ solution. Read the cell potential and record it on your report sheet.
7. Wipe off the metal electrodes with a paper towel before placing them aside for the next measurement. Be sure to keep alligator clips clean and non-contaminated with any solution throughout this process.
8. Repeat with the different metals and solutions indicated on the data sheet to complete the table, making sure to use a new salt bridge each time if a salt bridge is needed.
9. Write the equation for each electrode and determine the reaction quotient (Q), standard cell potential (E^o_{cell}), and cell potential for the given concentrations (E_{cell}) for each.
10. Dispose of liquids into the heavy metal waste container. Rinse well plates with both tap water and DI water.

Pre-Lab Questions

Prepare for lab by completing and understanding the answers to these questions. Refer to the Background or another resource, such as your textbook, if necessary.

1. What are the half-reactions occurring in a cell with a Fe/Fe(NO_3)$_3$ electrode and a Zn/Zn(NO_3)$_2$ electrode in the Zn| Zn^{2+}||Fe^{3+}|Fe cell?

2. In the Zn| Zn^{2+}||Fe^{3+}|Fe cell, what is the anode and what is the cathode?

3. If the cell potential for the Zn| Zn^{2+}||Fe^{3+}|Fe cell is 0.72 V, what is the reduction potential for the iron half-reaction?

4. Write the overall balanced equation and expression for Q for the Zn| Zn^{2+}||Fe^{3+}|Fe cell.

5. Which half-reaction in Table 1 is most likely to occur as an oxidation? Explain.

Voltaic Cells Report Sheet

Name ______________________ Section ______________________

Date ______________________ Instructor ______________________

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Cell	E_{cell} (meas)	Reaction quotient (Q)	E^{o}_{cell} (calc)	E_{cell} (calc)
Zn$\vert$ Zn^{2+}(1.0 M)$\Vert$Cu^{2+}(1.0 M)$\vert$Cu				
Zn$\vert$ Zn^{2+}(1.0 M)$\Vert$Cu^{2+}(0.10 M)$\vert$Cu				
Zn$\vert$ Zn^{2+}(0.10 M)$\Vert$Cu^{2+}(1.0 M)$\vert$Cu				
Zn$\vert$ Zn^{2+}(1.0 M)$\Vert$Pb^{2+}(1.0 M)$\vert$Pb				
Zn$\vert$ Zn^{2+}(1.0 M)$\Vert$Pb^{2+}(0.10 M)$\vert$Pb				
Zn$\vert$ Zn^{2+}(0.10 M)$\Vert$Pb^{2+}(1.0 M)$\vert$Pb				
Zn$\vert$ Zn^{2+}(1.0 M)$\Vert$Ag^{+}(0.10 M)$\vert$Ag				
Zn$\vert$ Zn^{2+}(0.10 M)$\Vert$Ag^{+}(0.10 M)$\vert$Ag				
Pb$\vert$ Pb^{2+}(1.0 M)$\Vert$Cu^{2+}(1.0 M)$\vert$Cu				
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Pb$\vert$ Pb^{2+}(0.10 M)$\Vert$Ag^{+}(0.10 M)$\vert$Ag				
Cu$\vert$ Cu^{2+}(1.0 M)$\Vert$Ag^{+}(0.10 M)$\vert$Ag				
Cu$\vert$ Cu^{2+}(0.10 M)$\Vert$Ag^{+}(0.10 M)$\vert$Ag				

Write the complete redox reaction for a Zn| Zn^{2+}||Cu^{2+}|Cu cell.

Write the complete redox reaction for a Zn| Zn^{2+}||Pb^{2+}|Pb cell.

Write the complete redox reaction for a Zn| Zn^{2+}||Ag^{+}|Ag cell.

Write the complete redox reaction for a Pb| Pb^{2+}||Cu^{2+}|Cu cell.

Write the complete redox reaction for a Pb| Pb^{2+}||Ag^{+}|Ag cell.

Write the complete redox reaction for a Cu| Cu^{2+}||Ag^{+}|Ag cell.