Voltaic Cells

Purpose

To set-up and measure voltages in voltaic cells and compare the actual values to theoretical values.

Learning Objectives

Determine the standard cell potentials for multiple cells by preparing voltaic cells and measuring the voltage.

Determine cell potential for non-standard cells by preparing voltaic cells and measuring the voltage.

Compare experimentally determined cell potentials to tabulated and calculated potentials.

Laboratory Skills

Use a Vernier voltmeter

Equipment

- Vernier voltmeter (mV precision)
- leads with alligator clips
- Microwells
- Strips of filter paper

Chemicals

- Zinc wire (3 cm x 0.5 cm)
- Copper wire (3 cm)
- iron wire (3 cm x 0.5 cm)
- Silver wire (3 cm)
- 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 $Fe(NO_3)_3$
- 0.1 M $Fe(NO_3)_3$
- 0.1 M $AgNO_3$
- 1.0 M KNO_3

Introduction

Electrochemistry

When electrons transfer between reaction components in a oxidation-reduction (redox) reaction, we can harness the motion of the electrons to create an electrical potential. Electrochemistry involves 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.

In order for the reaction to occur with the half-reactions physically separated, the electrons must flow through a wire connecting the electrodes, creating a current. If the redox reaction is a spontaneous reaction, the setup is called a voltaic or galvanic cell and 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.

Voltaic Cells

A cell with a spontaneous flow of electrons is known as a voltaic or galvanic cell, with the set-up as shown in Figure VC.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.

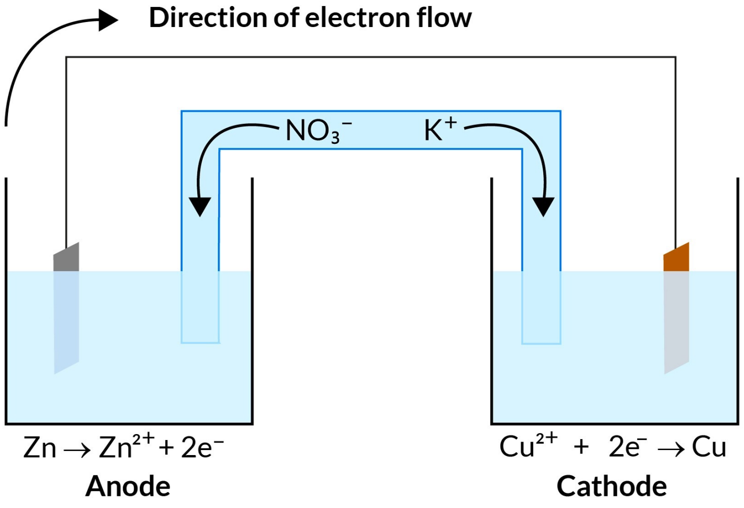

Figure VC.1: Schematic diagram of a voltaic cell with a zinc anode and a copper cathode

To complete an electrochemical 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 (half-cells). 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. In the diagram above, both the cathode and anode have excess KNO_3 dissolved in solution to act as spectator ions, and are shown moving across the salt bridge. The two half reactions from the two half-cells are shown in Reaction VC.1 and Reaction VC.2.

$$\mathrm{Zn(s)} \longrightarrow \mathrm{Zn^{2+}} + 2\,\mathrm{e^-} \qquad \text{(Reaction VC.1)}$$

$$\mathrm{Cu^{2+}} + 2\,\mathrm{e^-} \longrightarrow \mathrm{Cu(s)} \qquad \text{(Reaction VC.2)}$$

Reaction VC.3 shows the cell notation, which is a sort of shorthand for a voltaic cell. The cell notation is a set of symbols used to represent the components a cell.

$$\mathrm{Zn(s)\,|\,Zn^{2+}(1.0\,M)\,||\,Cu^{2+}(1.0\,M)\,|\,Cu(s)} \qquad \text{(Reaction VC.3)}$$

Cell notation follows the flow of electrons from left to right, starting with the anode. The reactant and product of the oxidation half-reaction appear with a vertical line separating the two phases (solid and aqueous). Next is a set of *double* vertical lines that indicate the separation between the two electrodes and represents the salt bridge. Finally, the reactant and products of the reduction half reaction of the cathode are shown, with a single vertical line separating the phases. Notice that the concentration of the aqueous solutions in included in both the anode and cathode portions of the cell notation.

Reduction Potentials

A voltaic cell creates a potential of usable electrical energy. Potential is measured in volts, where 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, the potentials of each half reaction must be considered. These half-reaction potentials are given as reduction potentials in Table VC.1.

Half-reaction potentials for reactions are typically given for reduction reactions only. The values shown are the reduction potentials as measured against the standard hydrogen electrode, which is assigned a value of 0 V. Any half-reactions with positive reduction potential will occur as written. The half-reactions with negative reduction potentials are more likely to occur in the reverse direction, as oxidation reactions.

Table VC.1: Select Standard Reduction Potentials

Reduction half reaction	**E^o_{red} (V)**
$Ag^+(aq) + e^- \longrightarrow Ag(s)$	+0.799
$Cu^{2+}(aq) + 2\,e^- \longrightarrow Cu(s)$	+0.337
$2\,H^+(aq) + 2\,e^- \longrightarrow H_2(g)$	0
$Fe^{3+}(aq) + 3\,e^- \longrightarrow Fe(s)$	−0.036
$Zn^{2+}(aq) + 2\,e^- \longrightarrow Zn(s)$	−0.763

The standard potential of a cell is calculated as the potential difference between the reduction potentials for the two half-reactions, as shown in Equation VC.1. Subtract the reduction potential at the anode where electrons start from the reduction potential at the cathode where electrons end up.

$$E^o_{\text{cell}} = E^o_{\text{red}}\,(\text{cathode}) - E^o_{\text{red}}\,(\text{anode}) \qquad \text{(Equation VC.1)}$$

Cell Conditions

The conditions of a cell also affect the cell potential. A **standard potential**, denoted by a degree sign, is one measured at a temperature of 298 K with all involved solutes present in one molar (1 M) concentrations. Potentials under non-standard conditions of temperature and concentration can be calculated using the Nernst equation, Equation VC.2.

$$E_{\text{cell}} = E^o_{\text{cell}} - \frac{RT}{nF}\ln Q \qquad \text{(Equation VC.2)}$$

The Nernst equation involves two constants, the gas constant, R (8.3145 J/K · mol, and Faraday's constant, F (96,485 C/mol). Temperature appears as T and must be in K (absolute temperature). Then, there are two components related to the balanced redox reaction. The n in the equation refers to the number of moles of electrons transferred in the balanced overall cell reaction. This number is the electron coefficient in the half-reactions when the overall reaction is balanced. Q is the reaction quotient, which is the same expression as equilibrium constant, but includes non-equilibrium concentrations. The reaction quotient includes reaction components that are solutions or gases and omits those that are pure solids and liquids. Thus, the value of the reaction quotient depends on the concentrations of the solutions used in the cell.

Procedure

Safety Precautions

Safety goggles are required!

Use the Vernier voltmeter and probes only as directed.

1. Connect the red (cathode) and black (anode) probes to the Vernier unit 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 1.0 M KNO_3 solution to act as salt bridges. **Make sure you know which well contains each solution.**
3. Take a piece of clean zinc wire and clip it to the black probe of the Vernier voltmeter. *Black* will always be the *anode.*
4. Place the zinc wire into the 1.0 M $Zn(NO_3)_2$ solution.
5. Clip a piece of clean copper wire to the red probe of the Vernier voltmeter.
6. Place the copper wire in the 1.0 M $Cu(NO_3)_2$ solution. Read and record the cell potential.
7. Wipe off the metal electrodes with a paper towel before placing them aside for the next measurement.

   > Keep alligator clips clean and dry throughout this process.

8. Repeat with the different metals and solutions indicated on the data sheet to complete the table. Use a new salt bridge each time a salt bridge is needed.
9. Write the equation for each electrode and determine the reaction quotient (Q), standard cell potential (E^ocell), 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.

Name: ______________________________

Section: __________ Date: ______________

Report Sheet:

Voltaic Cells

Report Table VC.1: Data and Calculations

Cell Notation	E_{cell} **(meas)**	**Reaction quotient (Q)**	E^{o}_{cell} (calc)	E_{cell} (calc)
Zn(s) | Zn^{2+}(1.0 M) || Cu^{2+}(1.0 M) | Cu(s)				
Zn(s) | Zn^{2+}(1.0 M) || Cu^{2+}(0.1 M) | Cu(s)				
Zn(s) | Zn^{2+}(0.1 M) || Cu^{2+}(1.0 M) | Cu(s)				
Zn(s) | Zn^{2+}(1.0 M) || Fe^{3+}(1.0 M) | Fe(s)				
Zn(s) | Zn^{2+}(1.0 M) || Fe^{3+}(0.1 M) | Fe(s)				
Zn(s) | Zn^{2+}(0.1 M) || Fe^{3+}(1.0 M) | Fe(s)				
Zn(s) | Zn^{2+}(1.0 M) || Ag^{+}(0.1 M) | Ag(s)				
Zn(s) | Zn^{2+}(0.1 M) || Ag^{+}(0.1 M) | Ag(s)				
Fe(s) | Fe^{3+}(1.0 M) || Cu^{2+}(1.0 M) | Cu(s)				
Fe(s) | Fe^{3+}(1.0 M) || Cu^{2+}(0.1 M) | Cu(s)				
Fe(s) | Fe^{3+}(0.1 M) || Cu^{2+}(1.0 M) | Cu(s)				
Fe(s) | Fe^{3+}(1.0 M) || Ag^{+}(0.1 M) | Ag(s)				
Fe(s) | Fe^{3+}(0.1 M) || Ag^{+}(0.1 M) | Ag(s)				
Cu(s) | Cu^{2+}(1.0 M) || Ag^{+}(0.1 M) | Ag(s)				
Cu(s) | Cu^{2+}(0.1 M) || Ag^{+}(0.1 M) | Ag(s)				

1. Write the complete (balanced) redox reaction for a $Zn(s) \mid Zn^{2+}(aq) \parallel Cu^{2+}(aq) \mid Cu(s)$ cell.

2. Write the complete (balanced) redox reaction for a $Zn(s) \mid Zn^{2+}(aq) \parallel Pb^{2+}(aq) \mid Pb(s)$ cell.

3. Write the complete (balanced) redox reaction for a $Zn(s) \mid Zn^{2+}(aq) \parallel Ag^{+}(aq) \mid Ag(s)$ cell.

4. Write the complete (balanced) redox reaction for a $Fe(s) \mid Fe^{3+}(aq) \parallel Cu^{2+}(aq) \mid Cu(s)$ cell.

5. Write the complete (balanced) redox reaction for a $Fe(s) \mid Fe^{3+}(aq) \parallel Ag^{+}(aq) \mid Ag(s)$ cell.

6. Write the complete (balanced) redox reaction for a $Cu(s) \mid Cu^{2+}(aq) \parallel Ag^{+}(aq) \mid Ag(s)$ cell.