

Chemistry 1225 Experiment 11 Voltaic Cells

Objective

I. Construct several small scale voltaic (galvanic) cells and measure their potentials. II. Calculate the standard reducing potential for each half-cell using the measured cell potential values. III. Use measured reduction potentials to construct an activity series for metals.

Introduction

If you who have amalgam fillings in your teeth you have probably experienced a painful sensation when you happened to bite onto a tiny sliver of aluminum foil from a chocolate bar or gum wrapper. That sensation was a low voltage electrical shock resulting from the oxidation-reduction reaction between the aluminum and the metallic amalgam components in an electrolytic environment of saliva.

An oxidation-reduction reaction can also be used to make a “potato clock” which consists of a battery made of one strip each of copper and zinc stuck into a raw potato (or any other juicy fruit or vegetable) and connected with wires to a digital clock.

The first battery was made in about 1800 by the Italian Physicist, Count Alessandro Volta. Volta’s battery consisted of a pile of alternating zinc and silver disks separated by paper soaked in salt water. He found that if the pile was tall enough he could detect a weak electrical shock if he touched the two ends of the pile. Although the methods have improved, Voltaic cells are still used to make the batteries that power our tools, toys, and appliances today. Voltaic cells are also called galvanic cells after Luigi Galvani, an Italian physician of the late 18th century who discovered that the muscles of dead frogs twitched when struck by static electricity. Learn more about Volta and Galvani at Wikipedia, Google, etc.

Voltaic cells

A voltaic cell consists of two half-cells electrically connected through an external circuit (voltmeter leads). Each half-cell consists of a single metal electrode immersed in a solution containing ions of that metal (e.g. Cu/Cu^{+2} as one half cell and Ag/Ag^+ as the other). It is important that the half-cells are separated, but ion flow must be allowed between the cells through a connecting salt bridge and electron flow is permitted through the external circuit. The salt bridge may be a sophisticated device or alternatively, may be a simple strip of filter paper soaked in a salt solution. As long there is an external circuit the electrons will flow from one cell to the other, oxidizing the metal in one cell and reducing the ions in the other cell. The salt bridge is necessary between the two half-cells to prevent a build-up of a positive charge in the cell where the metal is oxidized (the anode, loss of electrons, formation of cations) and a negative charge in the cell where the cations are reduced (the cathode, gain of electrons, formation of metal).

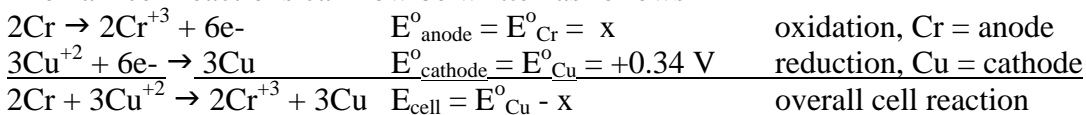
Each half-cell is characterized by a standard electrode reduction potential, E° . E° for each half-cell is determined at 298K by comparing the voltage produced by the half cell in question with the voltage produced by a standard hydrogen electrode. The values of E° for each half-cell allow one to determine which electrode will act as the anode (has a negative charge, acts as a source of electrons, and is the site of oxidation) and which

will act as the cathode (has a positive charge, acts as a sink for electrons, and is the site of reduction) in the galvanic cell. Using the equation, $E_{\text{cell}} = E^{\circ}_{\text{cathode}} - E^{\circ}_{\text{anode}}$, one can easily calculate the standard potential of the voltaic cell if one knows the value of E° for each half-cell. On the other hand, if the cell potential is measured and the standard reduction potential for one of the two half-cells is known, the reduction potential of the other half-cell can easily be calculated.

For this experiment you will construct a series of voltaic cells and use a voltmeter to measure their electrical potentials. The voltmeter displays the magnitude of the measured potential and indicates whether the electrode attached to the **red** lead is the anode or the cathode. A negative reading on the voltmeter indicates that the metal at the **red** lead is acting as an anode. That is, it is acting as a source of electrons and is being oxidized. A positive reading tells you that the metal at the **red** lead is acting as a cathode. That is, it is accepting electrons and is being reduced. The value of the cell potential, E_{cell} , is simply the absolute value of the measured voltage.

For example, assume that the cell potential was measured for a voltaic cell constructed from Cu/Cu^{+2} and Cr/Cr^{+3} half-cells and that E° for the Cu/Cu^{+2} half-cell is known to be 0.34 volts. If the red lead of the voltmeter was attached to the copper and the black lead was attached to chromium and the voltmeter display showed a value of +1.08 volts, one knows that because the red lead was on copper and the reading had a positive value, the copper electrode was the cathode and the chromium electrode was the anode. That is, copper ions were being reduced to copper metal and chromium metal was being oxidized to ions.

The half-cell reactions can now be written as follows



To find the reduction potential for chromium, $E^{\circ}(\text{Cr}/\text{Cr}^{+3})$, one needs to set up an equation to calculate the cell potential where

$$E_{\text{cell}} = E^{\circ}_{\text{cathode}} - E^{\circ}_{\text{anode}} \quad \text{Equation 1}$$

Since we know that $E_{\text{cell}} = E^{\circ}(\text{Cu}/\text{Cu}^{+2}) - E^{\circ}(\text{Cr}/\text{Cr}^{+3})$, that $E_{\text{cell}} = +1.08 \text{ V}$, and that $E^{\circ}(\text{Cu}/\text{Cu}^{+2}) = +0.34\text{V}$, it follows that

$$E^{\circ}(\text{Cr}/\text{Cr}^{+3}) = E^{\circ}(\text{Cu}/\text{Cu}^{+2}) - E_{\text{cell}} \quad \text{Equation 2}$$

Thus,

$$E^{\circ}(\text{Cr}/\text{Cr}^{+3}) = 0.34\text{V} - 1.08\text{V} = -0.74\text{V} \quad \text{Equation 3}$$

Part I. Reduction Potentials

In the first part of this experiment you will measure and tabulate the reduction potentials for a series of metals. Metals with a more positive reduction potential are stronger cathodes and the ions of that metal will receive electrons from any metal with a

more negative potential. You will then use the measured reduction potentials to construct an activity series for metals.

Part II. The Activity Series for Metals

In the second part of the experiment your activity series for metals will be verified by applying a solution of one metal ion onto the surface of a different solid metal. The more reactive metal (the reducing agent) will dissolve and the ions of the less reactive metal (oxidizing agent) will be reduced and the metal will precipitate.

For example, what will happen if a solution of copper ions [copper (II) sulfate or copper (II) nitrate] is placed on a piece of chromium metal? What would you expect to happen if the roles are reversed and chromium (III) nitrate (or sulfate) is placed on a piece of copper metal?

Solution:

The E° for the Cu/Cu^{+2} half-cell has a higher (more positive) value (0.34V) than the E° for the Cr/Cr^{+3} half cell (-0.74V). This means that copper ions are much stronger oxidizing agents than Cr^{+3} ions and that chromium metal is a stronger reducing agent than copper metal. In other words, chromium metal is more reactive than copper and as such chromium metal will reduce copper ions in solution. Thus, in a solution containing Cu^{+2} and Cr, copper ions will precipitate as copper metal and chromium metal will dissolve to form chromium (III) ions. In the solution containing Cr^{+3} and Cu, nothing should be observed since Cu will not reduce Cr^{+3} .

Procedure

Part I

1. Obtain a 9 cm filter paper and make a cell template according to Figure 1. Do not fold the paper and **use only pencil** to draw on it. Do not erase.
2. Write on each sector (using pencil only) the chemical symbol of each metal used in the experiment, as shown in figure 1.
3. Place the cutout filter paper onto the lid of a clean, dry plastic Petri dish.
4. Use tweezers to transfer the appropriate metal pieces onto their dedicated sections of the paper, Note: you are using steel nails instead of pure iron wire in this experiment.
5. Place 1-2 drops of the appropriate metal ion solution onto the paper at the edge of each metal piece. Make sure that the metal is contacting the solution. **Make sure that the ion solutions don't bleed enough to contact each other.**
6. Drop 1.0 M KNO_3 in the center of the paper to act as a salt bridge. Allow the KNO_3 solution to soak outward until it contacts all other wet areas.
7. Turn the voltmeter on and switch the dial to measure DC voltage. Set the range to 2V.

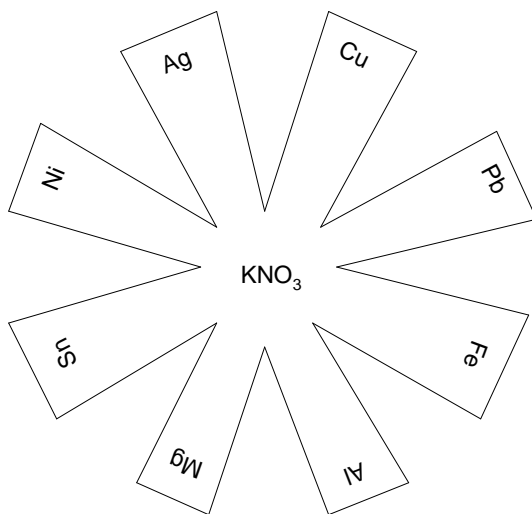


Figure 1. A Filter Paper Galvanic Cell

8. Place the red probe on Cu(s) and the black probe on Ag(s). Record the sign and value displayed.
9. Keep the red probe on Cu(s) and move the black probe to the next metal. Record the sign and value displayed.
10. While keeping the red probe on Cu(s) move the black probe to each consecutive metal. Record the sign and value for each Cu-metal combination.
11. Move the red probe to Ag(s). Keeping it on Ag(s), move the black probe around the circle making sure that you include the Cu(s) in this tour. **Note:** the measurements involving Al and Mg will be very unstable. Read and record the voltage value as quickly as possible after you touch the probe to the metal.

Part II

Place a piece of zinc (use the piece from part I) on a Petri dish and add one drop of 1.0 CuSO₄ solution onto the metal surface. Observe and record any changes that occur. Obtain a piece of copper foil and place a drop of ZnSO₄ onto it. Observe and record your observations.

Report

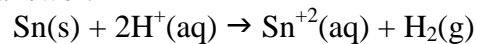
Calculations: calculate the reducing potentials for each of the half-cells tested using Cu/Cu⁺² as a reference ($E^{\circ} = +0.34\text{V}$) and your measured reduction potentials for the different metals. Then using the calculated value for the Ag/Ag⁺ half-cell, repeat the

calculations for each of the other metals (remember that you measured each metal against Ag/Ag^+). Calculate the average E° value for each metal.

Results and discussion: create an activity table showing your standard reduction potential values in order of increasing reduction potential. Include the literature values for each metal as a reference. Discuss the reasons for any differences between the experimentally determined values and the literature values. Include in your discussion an analysis of part II of the experiment describing what happened for each part of the experiment and why. Also include the half-cell reactions for part II and the overall reaction equation.

Conclusion: summarize the experiment, its goals, and your results in two to four sentences.

1. Use the activity series from your text to determine whether tin will dissolve in a 1.0M acid solution (a solution of H^+ ions) according to the following equation. Explain your answer.



2. Would a battery constructed of tin (Sn/Sn^{+2}) and lead (Pb/Pb^{+2}) be of much value? Why or why not? Hint: calculate the reduction potential for the battery.
3. Martin Smythe constructed a potato battery by inserting strips of copper and zinc into a potato. The metal strips were attached to the electrical leads of a clock. What is the voltage produced by the battery?
4. What is the purpose of the potato in question 3?

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Based on your experimentally determined activity series answer the following questions.

1. Which metal is the strongest oxidizing agent in your activity table? Which metal in your table is the strongest reducing agent? Explain.

2. If you wanted to design a battery from relatively common, inexpensive, and not too toxic materials, that produced a voltage of about 1V, which cell reaction would you use?

3. A. Will metallic magnesium precipitate if a piece of tin is added to a 1.0 M solution of magnesium sulfate? Explain.

B. Will a piece of copper dissolve in a 1.0 M solution of $\text{Pb}(\text{NO}_3)_2$? Explain.