

Electrochemical Cells

Introduction

The use of electrochemical cells to convert the Gibbs energy stored in the constituent half-reactions into electrical work is of enormous industrial as well as fundamental significance. We have all used batteries and these are simply galvanic cells that are similar to those you will be constructing in this experiment. In the laboratory, a typical electrochemical cell has the following general construction:

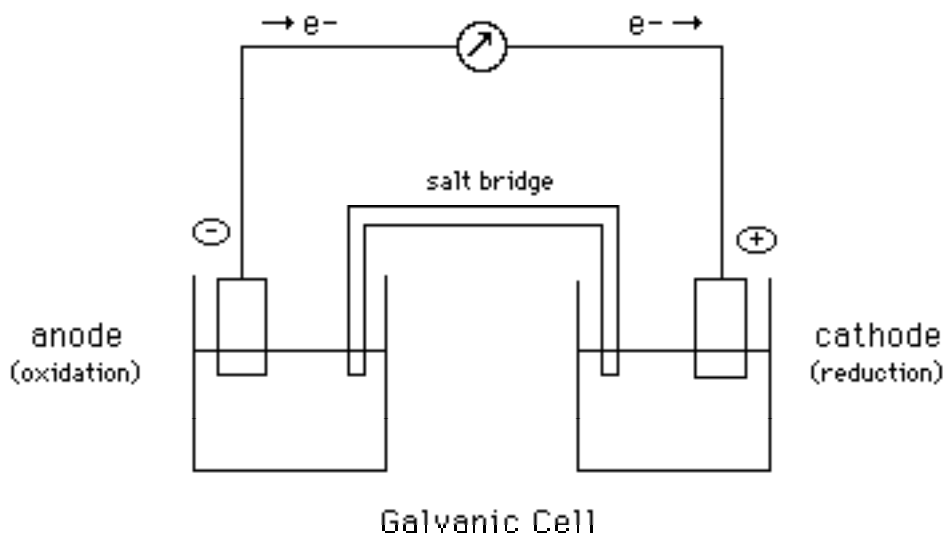


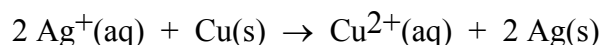
Figure 1: A Schematic of a Galvanic Cell

In Figure 1, there are two electrode-compartments, each of which contains an electrode and the constituents of the half-reaction. In many instances, the electrode is actually one of the chemical components of the half-reaction. For example, the copper electrode is involved directly in the half reaction, $\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$. In other cases, however, the electrode does not participate in the chemistry of the half-reaction, but merely provides an inert conducting surface on which the electron exchange occurs. For example, when the half reaction, $\text{Fe}^{3+} + \text{e}^- \rightarrow \text{Fe}^{2+}$ is studied a Pt or C electrode rather than an iron electrode is used. This is because on an iron wire, both $\text{Fe}^{3+} + 3\text{e}^- \rightarrow \text{Fe}(\text{s})$ and $\text{Fe}^{2+} + 2\text{e}^- \rightarrow \text{Fe}(\text{s})$ could occur, rather than the $\text{Fe}^{3+}/\text{Fe}^{2+}$ reaction that is of interest.

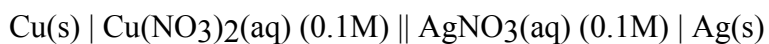
The two electrode compartment must be separated by a barrier that permits ions to migrate inside the cell, but does not allow the contents of the chambers to mix. A glass tube that is filled with a gel saturated with a strong electrolyte such as KNO_3 is often used. In this case, the K^+ and NO_3^- ions create a "salt bridge" for the electrochemical cell.

Electrochemical cells have both a magnitude for the measured voltage and a polarity. An electrode that is positive, relatively speaking, must be deficient in electrons, and hence a reduction must be taking place at that electrode. Conversely, an electrode that appears negative has a surplus of electrons. Hence, electrons are being released into it by an oxidative half-reaction. By definition, oxidation occurs at an anode. Hence, the (-) pole of an electrochemical cell is necessarily its anode, and the (+) pole is its cathode (where the reduction occurs).

The directions for this experiment exploit the fact that electrochemical cells can be described very efficiently by using conventional "cell diagrams." A possible diagram for a galvanic cell that employs the net ionic reaction,



is



The corresponding measured conventional cell voltage is roughly +0.4 V.

The cell diagram contains the information necessary to construct the electrochemical cell. The cell diagram has a "right hand" electrode (here a piece of silver wire) and a "left hand" electrode (here a piece of copper wire). A single vertical line means that the two species flanking it have different phases. In this case, the two phases are the solid wires and the liquid solutions whose concentrations are specified inside the parentheses. The double vertical line implies some type of chemically inert salt bridge connecting the two compartments in question.

The algebraic sign of a conventionally diagrammed cell voltage is the experimentally determined polarity of the "right hand" electrode. If the **cell voltage is positive**, electrons must be being removed from the right hand electrode (cathode) by a reductive half-reaction (remember **R**eduction on the **R**ight). *There is, however, nothing that absolutely requires the voltage of diagrammed cell to be positive.* Negative conventional cell voltages are allowed; in fact, one or more of the cell voltages you will be measuring in this experiment will turn out to be negative. *All this means is that oxidation is what really takes place at the right hand electrode (anode) of the cell as diagrammed. This indicates that the cell diagram is written in reverse of the spontaneous direction of the cell.* Remember that a correct cell diagram of a spontaneous redox reaction has the anode written on the left of the double vertical line and the cathode written to the right of the double vertical line.

In this experiment you will construct and measure the voltage of electrochemical cells that involve the half-reactions (in alphabetical order): $\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$, $\text{Fe}^{3+} + \text{e}^- \rightarrow \text{Fe}^{2+}$, $\text{Pb}^{2+} + 2\text{e}^- \rightarrow \text{Pb}$, $\text{Zn}^{2+} + 2\text{e}^- \rightarrow \text{Zn}$. From your results you will be able to determine the relative positions of these half-reactions in a Table of Standard Potentials. You will also confirm that the potentials of half-reactions are concentration dependent. As you know, in order to calculate the voltages of non-standard cells (not 1 M and 1 atm), you need to use the Nernst equation. A useful form of the Nernst equation is equation 21.18 in section 21-4 on page 839 in your textbook. **Read Chapter 21 in**

your Petrucci textbook, 8th ed., as pre-laboratory preparation before beginning this experiment.

There is one last but important note. This experiment is quite easy and short in terms of the actual collection of data. However, the write-up is difficult and requires a good understanding of electrochemistry. Students in the past have found it very helpful if they take the time to do each step slowly and try to understand what is happening. They also have said that is critical to begin the write-up while they are in the laboratory in order to repeat steps that have become confusing or to get assistance from the TA or fellow students. You are strongly encouraged to follow this advice!

Preparation for Next Lab. Each student needs to dry a sample of solid calcium carbonate and a solid unknown for the next laboratory (EDTA Titrations).

- 1) Half fill one vial with pure calcium carbonate.
- 2) Obtain a vial of the unknown calcium carbonate mixture from your TA. Place an X on this vial with a grease pencil to distinguish it from the pure calcium carbonate vial.
- 3) Place the uncapped vials in a beaker. Write your name on the white frosted area of the beaker and place it in the oven. Do NOT use PAPER labels on your vials or beaker. Cover the beaker with a watch glass.
- 4) Dry the samples in the oven for 1.5 hours. Do not adjust the temperature on the oven. The temperature on the oven has been preset and will heat to the correct temperature when the door remains closed.
- 5) After removing your samples from the oven, let them cool until they are warm but safe to handle.
- 6) After the samples have cooled carefully, place beaker containing the uncapped vials in the desiccator until needed. Be careful not to touch the vials with your fingers.

Make sure that you label the vials!

Safety, Clean-up and Special instructions:

Wear your Goggles and gloves throughout this experiment.

Dispose of all solutions on your spot plate by holding it vertically above the mouth of the funnel in the neck of the metal waste disposal container and rinsing it with a brisk stream from your wash bottle. The metal waste containers are in the fumehoods. Wash the spot plate more thoroughly under the deionized water faucet and dry it for later use.

Excess solutions containing lead and copper need to be disposed of in the metal waste containers while the excess solutions containing only iron may go down the drain with copious amounts of water.

Treat the voltmeters with care and respect as they are expensive.

Procedure

You will work in pairs on this experiment.

The actual data analyses and the written reports must be done entirely independently of your lab partner or other students. Make sure that you avoid unauthorized collaboration and plagiarism. All suspected violations of the Code of Academic Conduct will be referred to Student Judicial Affairs.

Part I. Constructing a Table of Standard Reduction Potentials

In this part of the experiment you will be measuring the voltages of the several galvanic cells diagrammed below. In other words, you will be using the "V" scales of the meters (which you will find at the front of the laboratory.) Note that the meter display has both an algebraic sign and a magnitude. THROUGHOUT this experiment you should clip the "red" wire coming out of the meter to the "right hand" electrode of the cell in question. If the meter shows a "plus" voltage, reduction takes place at that electrode. If the displayed voltage is negative, oxidation is what really occurs at the "right hand" electrode of the cell as diagrammed.

In most cases the cell electrolytes for Part I of the experiment will be 0.1 M stock solutions of the nitrate salts of the various cations. For the cell containing a mixture that is 0.1 M in both Fe^{2+} and Fe^{3+} you should mix together 1 mL volumes of the 0.2 M stock solutions of the respective sulfate or nitrate salts in a clean, dry test tube. Note that the Fe^{2+} stock solution contains sulfuric acid and some iron nails. This is to insure that any ferric ion that might get produced by air oxidation gets re-reduced to the 2+ oxidation state. Because oxygen does attack Fe^{2+} , you should measure the voltages of cells containing the ferrous ion as soon as possible after the solution in question has been prepared. These cell voltages will probably change over time.

The electrode "compartments" are simply the wells in a "spot plate." A well will hold about 1 mL of solution and they can be filled very easily using disposable polyethylene pipets. The "salt bridge" is nothing more than a well containing 0.1 M sodium nitrate solution which gets connected to two or more other wells with short (ca. 3 cm) lengths of cotton string that you have pre-saturated in a watch glass containing 0.1 M NaNO₃ solution. Note: You CANNOT connect the anode and cathode wells directly with the NaNO₃ saturated string. You MUST connect both the anode and the cathode to the well containing the 0.1 M NaNO₃ solution with the string.

Use the **plastic blue tongs** that are available in the laboratory to manipulate the strings. The strings can actually function as siphons so you should always fill all the desired wells in the spot plate first and then put the conducting strings in place.

Assemble and measure the voltages of the following conventional cells:

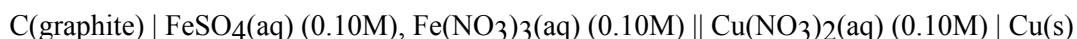
Cell #1



Cell #2



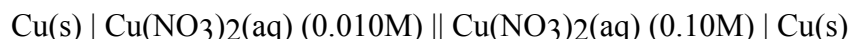
Cell #3



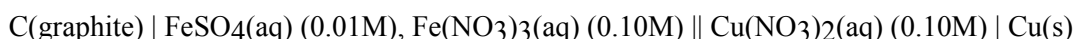
Part II. The Concentration Dependence of Half Cell Potentials

Assemble and measure the voltages of the following conventional cells:

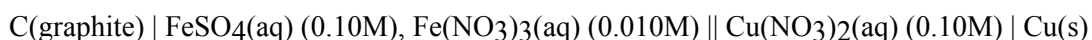
Cell #4



Cell #5



Cell #6



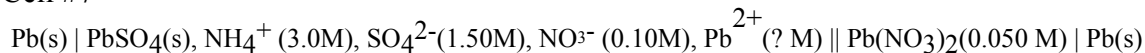
The "left hand" compartment electrolyte solutions are prepared in the first by diluting either the 0.10 M or the 0.20 M stock solutions by a factor of 10. This involves adding 1 mL of the 0.20 M stock solution to 9 mL of deionized water. Then, second, by mixing the diluted solutions with the undiluted other ones in equal volumes.

Clean up: Dispose of the excess diluted copper solution in the waste container. The iron solutions can be washed down the drain.

Part III. Estimating the Solubility Product of Lead(II) Sulfate

For this part of the experiment prepare a 0.050 M $\text{Pb}(\text{NO}_3)_2$ solution by diluting the stock 0.1 M $\text{Pb}(\text{NO}_3)_2$ solution by a factor of 2 in one test tube. In another test tube mix equal volumes of 0.10 M $\text{Pb}(\text{NO}_3)_2$ and 3.0 M ammonium sulfate. Allow about two minutes for the complete precipitation of PbSO_4 . Assemble and measure the voltage of the cell:

Cell #7



(This is experimentally very easy to do: fill one well of your spot plate with the diluted lead nitrate solution, fill an adjacent one with a slurry of the contents of the test tube containing the precipitated lead sulfate, connect them with salt bridge strings to a well full of the sodium nitrate solution, attach two Pb electrodes to the wires from the meter, and measure the cell voltage.)

Clean up. Dispose of all the lead-containing solutions in the proper waste container found in the fumehood.

Data Analysis

Post Laboratory Exercises will guide you through these calculations.

Part I.

For every conventional galvanic cell diagram there is a corresponding (chemical) cell reaction, as illustrated by the example in the Introduction section of this experiment. According to the Nernst equation the conventional cell voltage of such a cell can be written in general as

$$E_{\text{cell}} = (E^{\circ}_{\text{cathode}} - E^{\circ}_{\text{anode}}) - [0.0257 / n] \ln Q$$

where $E^{\circ}_{\text{cathode}}$ is the standard reduction potential of the cathode half cell reaction and E°_{anode} is the standard reduction potential of the anode half cell reaction. Q is the reaction quotient, for the overall cell reaction, using the concentrations given in the cell diagram. Note that at equilibrium when $Q = K$ then E_{cell} is equal to zero.

For the cell discussed in the introduction this equation becomes:

$$\begin{aligned} E_{\text{cell}} &= (E^{\circ}_{\text{Ag}^+/\text{Ag}} - E^{\circ}_{\text{Cu}^{++}/\text{Cu}}) - 0.0129 \ln[(0.10\text{M}) / (0.10\text{M})^2] \\ &= (0.799 - 0.337) - 0.0297 \\ &= + 0.432 \text{ V} \end{aligned}$$

The numerical values of the standard reduction potentials used in this example were taken from a table in the Appendix of your textbook.

In part I you measured the voltages of three cells, each containing a different metal, against a common Cu(II)/Cu reference electrode. The standard reduction potential of Cu(II)/Cu(s) redox couple is known to be 0.34V. This value is used as the "known" standard reduction potential in the following calculations.

For each of the three cells whose voltages you measured in Part I of this experiment write the corresponding redox reaction in the spontaneous direction and calculate the value of the standard cell potential, ($E^{\circ}_{\text{cathode}} - E^{\circ}_{\text{anode}}$) using your measured cell potentials and the Nernst equation. Using the standard cell potential you just calculated and the known value of the standard reduction potential of the Cu(II)/Cu(s) redox couple, determine the standard reduction potential of each of the noncopper metal half cells. You should also compare your findings with the accepted values.

Part II.

(A) Write the chemical reaction in the spontaneous direction for the copper concentration cells--the first cell you examined in Part II of the experiment. Calculate the theoretical value of the potentials for these copper concentration cells using the Nernst equation. Compare the measured and theoretical values for these cell potentials. Your agreement will be only semi-quantitative.

(B) Altogether, you have measured the voltages of three cells with different relative concentrations of Fe^{2+} and Fe^{3+} ions against a common Cu^{2+}/Cu reference electrode. If you write out the Nernst equation for these three cells expanding the logarithmic terms you will find that in each case the cell voltage can be written as a sum of four terms. [Hint: $\ln(a b / c) = (\ln a) + (\ln b) - (\ln c)$] Three of them are invariant and include the E° 's of the iron and copper half reactions along with the term that takes the log of the copper ion concentration into account. The fourth term, however, varies in value and involves only the log of the ratio of the iron species concentrations, and must therefore be responsible for the voltage changes you observed in the laboratory. Calculate the theoretical cell voltage DIFFERENCES for the three concentration cells involving the copper and iron and compare your experimental results with these predictions. Again you should expect only semi-quantitative agreement. In separate calculations, find the cell voltage difference between cell #3 in Part I and each of the iron containing cells of Part II.

Part III.

In this experiment what you have really done is prepare a concentration cell with drastically different concentrations of the Pb^{2+} ion in the two halves of the cell. Use the Nernst equation and the measured cell voltage to estimate the Pb^{2+} ion concentration in the compartment containing the lead sulfate slurry. The solubility product of lead sulfate can then be estimated by multiplying this result by the concentration of the sulfate ion in that compartment. Your result should be within about a factor of ten of the tabulated value for this constant.

Post-Laboratory Exercise Questions

Part I analysis. The following series of questions and calculations will lead you through the calculation of the standard reduction potentials of the half-cells, $\text{Zn}^{2+}/\text{Zn(s)}$, $\text{Pb}^{2+}/\text{Pb(s)}$, & $\text{Fe}^{3+}/\text{Fe}^{2+}$.

Part I Analysis. Given the Nernst equation, $E_{\text{cell}} = (E^{\circ}_{\text{cathode}} - E^{\circ}_{\text{anode}}) - [0.0257/n]\ln Q$, and the equation, $E^{\circ}_{\text{cell}} = E^{\circ}_{\text{cathode}} - E^{\circ}_{\text{anode}}$, substitute E°_{cell} into the Nernst equation and solve for E°_{cell} . Which of the following would be the correct expression for E°_{cell} ? 3 choices will be given.

Part I analysis. For the cell, $\text{Zn(s)} | \text{Zn(NO}_3)_2(0.10\text{M}) || \text{Cu(NO}_3)_2(0.10\text{M}) | \text{Cu(s)}$ which of the following is the correct spontaneous overall cell reaction: 3 choices will be given

Part I analysis. Please enter the value of E_{cell} you measured in volts for the spontaneous cell reaction for $\text{Zn(s)} | \text{Zn(NO}_3)_2(0.10\text{M}) || \text{Cu(NO}_3)_2(0.10\text{M}) | \text{Cu(s)}$.

Part I analysis. For the cell, $\text{Zn(s)} | \text{Zn(NO}_3)_2(0.10\text{M}) || \text{Cu(NO}_3)_2(0.10\text{M}) | \text{Cu(s)}$ which of the following is the correct expression for Q , the reaction quotient in the Nernst equation, for the spontaneous overall cell reaction: 3 choices will be given

Part I analysis. Using your cell concentrations and the reaction quotient expression, calculate and enter the value of Q for the Nernst equation for this cell $\text{Zn(s)} | \text{Zn(NO}_3)_2(0.10\text{M}) || \text{Cu(NO}_3)_2(0.10\text{M}) | \text{Cu(s)}$.

Part I analysis. Please enter the value of $\ln Q$ that appears in the Nernst equation for this cell, $\text{Zn(s)} | \text{Zn(NO}_3)_2(0.10\text{M}) || \text{Cu(NO}_3)_2(0.10\text{M}) | \text{Cu(s)}$.

Part I analysis. For the cell, $\text{Zn(s)} | \text{Zn(NO}_3)_2(0.10\text{M}) || \text{Cu(NO}_3)_2(0.10\text{M}) | \text{Cu(s)}$, put all the terms together that enter into the equation for E°_{cell} and enter the value you calculate for it.

Part I analysis. Please enter the accepted value from the table in your text for the standard electrode potential, E° , for the half-cell reaction for copper: $\text{Cu}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Cu(s)}$ in volts. You will use this value as a reference to calculate the standard electrode potentials for the other half-reactions involved in the cells you measured.

Part I analysis. Using the equation $E^{\circ}_{\text{cell}} = E^{\circ}_{\text{cathode}} - E^{\circ}_{\text{anode}}$, and the reference value of E° for Cu(II)/Cu(s) for the role that copper plays in this overall cell reaction, determine the value of the standard potential for the half-reaction $\text{Zn}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Zn(s)}$ and enter it here.

Part I analysis. For cell $\text{Pb(s)} | \text{Pb(NO}_3)_2(0.10\text{M}) || \text{Cu(NO}_3)_2(0.10\text{M}) | \text{Cu(s)}$, which of the following is the correct spontaneous overall cell reaction: 3 choices will be given

Part I analysis. Please enter the value of E_{cell} you measured in volts for the spontaneous cell reaction for For cell $\text{Pb(s)} | \text{Pb(NO}_3)_2(0.10\text{M}) || \text{Cu(NO}_3)_2(0.10\text{M}) | \text{Cu(s)}$.

Part I analysis. For the cell, $\text{Pb(s)} | \text{Pb(NO}_3)_2(0.10\text{M}) || \text{Cu(NO}_3)_2(0.10\text{M}) | \text{Cu(s)}$ which of the following is the correct expression for Q , the reaction quotient in the Nernst equation, for the spontaneous overall cell reaction: 3 choices will be given

Part I analysis. Using your cell concentrations and the reaction quotient expression, calculate and enter the value of Q for the Nernst equation for this cell $\text{Pb(s)} | \text{Pb(NO}_3)_2(0.10\text{M}) || \text{Cu(NO}_3)_2(0.10\text{M}) | \text{Cu(s)}$

Part I analysis. Please enter the value of $\ln Q$ that appears in the Nernst equation for this cell, $\text{Pb(s)} | \text{Pb(NO}_3)_2(0.10\text{M}) || \text{Cu(NO}_3)_2(0.10\text{M}) | \text{Cu(s)}$

Part I analysis. For the cell, $\text{Pb(s)} | \text{Pb(NO}_3)_2(0.10\text{M}) || \text{Cu(NO}_3)_2(0.10\text{M}) | \text{Cu(s)}$, put all the terms together that enter into the equation for E°_{cell} and enter the value you calculate for it.

Part I analysis. Using the equation $E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$, and the reference value of E° for Cu(II)/Cu(s) for the role that copper plays in this overall cell reaction, determine the value of the standard potential for the half-reaction $\text{Pb}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Pb(s)}$ and enter it here.

Part I Analysis. For cell $\text{C(gr)} | \text{Fe(NO}_3)_3(0.10\text{M}), \text{FeSO}_4(0.10\text{M}) || \text{Cu(NO}_3)_2(0.10\text{M}) | \text{Cu(s)}$, which of the following is the correct spontaneous overall cell reaction: 3 choices will be given

Part I analysis. Please enter the value of E_{cell} you measured in volts for the spontaneous cell reaction for cell $\text{C(gr)} | \text{Fe(NO}_3)_3(0.10\text{M}), \text{FeSO}_4(0.10\text{M}) || \text{Cu(NO}_3)_2(0.10\text{M}) | \text{Cu(s)}$.

Part I analysis. For the cell, $\text{C(gr)} | \text{Fe(NO}_3)_3(0.10\text{M}), \text{FeSO}_4(0.10\text{M}) || \text{Cu(NO}_3)_2(0.10\text{M}) | \text{Cu(s)}$, which of the following is the correct expression for Q , the reaction quotient in the Nernst equation, for the spontaneous overall cell reaction: 3 choices will be given

Part I analysis. Using your cell concentrations and the reaction quotient expression, calculate and enter the value of Q for the Nernst equation for this cell $\text{C(gr)} | \text{Fe(NO}_3)_3(0.10\text{M}), \text{FeSO}_4(0.10\text{M}) || \text{Cu(NO}_3)_2(0.10\text{M}) | \text{Cu(s)}$.

Part I analysis. Please enter the value of $\ln Q$ that appears in the Nernst equation for this cell, $\text{C(gr)} | \text{Fe(NO}_3)_3(0.10\text{M}), \text{FeSO}_4(0.10\text{M}) || \text{Cu(NO}_3)_2(0.10\text{M}) | \text{Cu(s)}$.

Part I analysis. For the cell, $\text{C(gr)} | \text{Fe(NO}_3)_3(0.10\text{M}), \text{FeSO}_4(0.10\text{M}) || \text{Cu(NO}_3)_2(0.10\text{M}) | \text{Cu(s)}$, put all the terms together that enter into the equation for E°_{cell} and enter the value you calculate for it.

Part I analysis. Using the equation $E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$, and the reference value of E° for Cu(II)/Cu(s) for the role that copper plays in this overall cell reaction, determine the value of the standard potential for the half-reaction $\text{Fe}^{3+}(\text{aq}) + \text{e}^- \rightarrow \text{Fe}^{2+}(\text{aq})$ and enter it here.

Part II analysis. Using the facts that a spontaneous cell reaction gives a positive measured voltage when the red lead of the voltmeter is connected to the cathode, and that reduction takes place at the cathode while oxidation takes place at the anode, which of the following is the spontaneous reaction for the copper concentration cell you measured in Part II of the electrochemistry experiment? 3 choices will be given

Part II analysis. Because the standard potentials are the same for both the anode and the cathode reactions for a concentration cell, the Nernst equation for a concentration cell becomes simply $E_{\text{cell}} = -(0.0257/n)\ln Q$, where Q is to be calculated for the spontaneous cell reaction. What is the value of Q for the copper concentration cell?

Part II analysis. $E_{\text{cell}} = -(0.0257/n)\ln Q$ Now, using your value for Q calculate E_{cell} from the Nernst equation for the copper concentration cell, and enter it here.

For Part II of the electrochemistry experiment, you measured the potential of some concentration cells. The laboratory manual asked you to construct the cells

a) $\text{Cu(s)}|\text{Cu(NO}_3)_2(0.010\text{M}) || \text{Cu(NO}_3)_2(0.10\text{M})|\text{Cu(s)}$

b) C(graphite) | FeSO₄(0.010M), Fe(NO₃)₃(0.10M) || Cu(NO₃)₂(0.10M) | Cu(s)

c) C(graphite) | FeSO₄(0.010M), Fe(NO₃)₃(0.10M) || Cu(NO₃)₂(0.10M) | Cu(s)

When the red lead for the voltmeter was connected to the right-hand electrode of the cells as diagrammed above, which of these cells showed a positive algebraic sign for the measured voltage?

Part II analysis. When a measured cell voltage is negative, the electrode to which the red lead is attached is actually the anode rather than the cathode. The anode is the electrode at which oxidation takes place, so it is surrendering electrons to the external circuit. For the two cells involving the iron and copper species, the Cu(II)/Cu(s) couple is the anode. Which of the following half-cell reactions is taking place at the Cu electrode in the anode compartment of these two cells? 3 choices will be given

Part II analysis. You will now calculate the theoretical cell voltage differences for the three concentration cells involving the copper and iron, cells #3, #5, and #6, and compare your experimental results with these predictions.

Part II analysis. For the overall cell reaction in cells #3, #5, & #6, $\text{Cu(s)} + 2\text{Fe}^{3+}(\text{aq}) \rightarrow 2\text{Fe}^{2+}(\text{aq}) + \text{Cu}^{2+}(\text{aq})$, which of the following is the correct expression for Q, the reaction quotient in the Nernst equation: 3 choices will be given

Part II analysis. The Nernst equation for the cell potential of cells #3, #5, & #6, is $E_{\text{cell}} = E^{\circ}_{\text{cell}} - (0.0257/2)\ln(Q_{\text{previous question}})$. What is the expression, then, for the cell voltage difference between cells #5 and #3, $E_{\text{cell}\#5} - E_{\text{cell}\#3}$? 3 choices will be given

Using the ion concentration values in cells #3 & #5 in the expression obtained from the previous question, calculate the theoretical cell voltage difference between cells #5 and #3.

Using the ion concentration values in cells #3 & #6 and the expression for $E_{\text{cell}\#6} - E_{\text{cell}\#3}$, calculate the theoretical cell voltage difference between cells #6 and #3.

Part III analysis. You will now calculate the solubility product for PbSO₄ by using your measured potential for cell #7 in Part III of the laboratory.

Part III analysis. Please enter the value of E_{cell} you measured in volts for the spontaneous cell reaction for cell $\text{Pb(s)} | \text{PbSO}_4(\text{s}), \text{NH}_4^+ (3.0\text{M}), \text{SO}_4^{2-}(1.50\text{M}), \text{NO}_3^- (0.10\text{M}), \text{Pb}^{2+} (? \text{M}) || \text{Pb}(\text{NO}_3)_2(0.050 \text{M}) | \text{Pb(s)}$.

Part III analysis. Because the standard potentials are the same for both the anode and the cathode reactions for a concentration cell, the Nernst equation for a concentration cell becomes simply $E_{\text{cell}} = -(0.0257/n)\ln Q$, where Q is the expression for the spontaneous cell reaction. For the concentration cell, $\text{Pb(s)} | \text{PbSO}_4(\text{s}), \text{NH}_4^+ (3.0\text{M}), \text{SO}_4^{2-}(1.50\text{M}), \text{NO}_3^- (0.10\text{M}), \text{Pb}^{2+} (? \text{M}) || \text{Pb}(\text{NO}_3)_2(0.050 \text{M}) | \text{Pb(s)}$ which of the following is the correct expression for Q for the spontaneous overall cell reaction, $\text{Pb}^{2+}(0.050\text{M}) \rightarrow \text{Pb}^{2+}(?)$: 4 choices will be given

Part III analysis. For the concentration cell, $\text{Pb(s)} | \text{PbSO}_4(\text{s}), \text{NH}_4^+ (3.0\text{M}), \text{SO}_4^{2-}(1.50\text{M}), \text{NO}_3^- (0.10\text{M}), \text{Pb}^{2+} (? \text{M}) || \text{Pb}(\text{NO}_3)_2(0.050 \text{M}) | \text{Pb(s)}$, $E_{\text{cell}} = -(0.0257/2)\ln Q_{\text{previous question}}$, what is the expression for x in terms of E_{cell} ? 4 choices will be given.

Part III analysis. Using your measured cell potential for $\text{Pb(s)} | \text{PbSO}_4(\text{s}), \text{NH}_4^+ (3.0\text{M}), \text{SO}_4^{2-}(1.50\text{M}), \text{NO}_3^- (0.10\text{M}), \text{Pb}^{2+} (? \text{M}) || \text{Pb}(\text{NO}_3)_2(0.050 \text{M}) | \text{Pb(s)}$ in your expression for x, calculate the molarity of Pb^{2+} in the anode compartment.

Part III analysis. Which of the following is the correct solubility product expression, K_{sp} , for $PbSO_4$? 5 choices will be given

Part III analysis. Using the expression, $K_{sp} = [Pb^{2+}][SO_4^{2-}]$, the concentration of sulfate in the anode compartment, and determined value Pb^{2+} concentration in the anode compartment, calculate K_{sp} . Please enter your molarity in decimal form only. Do not use scientific notation.