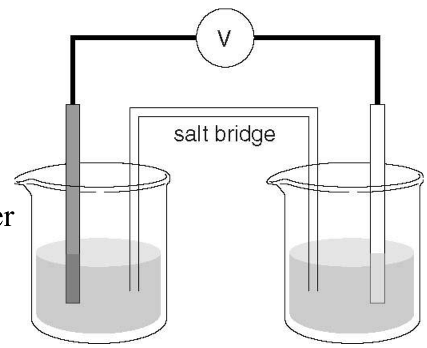
#25 Electrochemistry

Purpose: An oxidation-reduction reaction is observed. A voltaic cell is assembled and its potential measured. An electrolytic cell is used to copper plate an object.

Introduction
A brief description of voltaic and electrolytic cells is given below.
Voltaic cell
In a voltaic cell an oxidation-reduction (redox) reaction produces electrical energy by combining two half cells. The cell shown is made from copper and zinc half-cells. In the copper half-cell a strip of Cu metal, the Cu electrode, is immersed in a copper solution; in the zinc half cell a Zn electrode is in the zinc solution.



The copper is the cathode and zinc is the anode. Oxidation or *loss* of electrons takes place at the zinc anode:

$$Zn - 2e^- \rightarrow Zn^{2+}$$

Reduction or *gain* of electrons takes place at the copper cathode:

$$Cu^{2+} + 2e^- \rightarrow Cu$$

To find the cell voltage we combine their standard "reduction" potentials, E^o, listed in the table. For the cathode, use the reduction potential as listed. For the anode change the sign of the standard reduction potential listed. Then add the two.

Half Reaction	Reduction Potential E^o (V)
$Ag^+ + e^- \rightarrow Ag$	0.80
$Cu^{2+} + 2e^- \rightarrow Cu$	0.34
$Sn^{2+} + 2e^- \rightarrow Sn$	-0.14
$Zn^{2+} + 2e^- \rightarrow Zn$	-0.76

For the copper/zinc cell, 0.34 V is the reduction potential for the copper cathode. For the zinc anode, the sign of the potential (-0.76) is changed giving 0.76. The predicted voltage is:

$$0.34 + 0.76 = 1.10 \text{ V}$$

Electrolytic Cell
In an electrolytic cell electrical energy is supplied by a DC source (battery). The electrodes can be made of inert material since they are surfaces on which the redox reactions occur and do not participate in the reaction. Electrolysis can be used for electroplating. The object to be plated is the cathode where reduction takes place. The solution contains the metal to be plated. For copper plating, Cu^{2+} ions are reduced to copper metal:

$$Cu^{2+} + 2e^- \rightarrow Cu$$

Electrolysis is also used for decomposing compounds into their elements, for example NaCl into sodium metal and chlorine gas or H_2O into hydrogen and oxygen gas.

Safety and Waste Disposal Wear gloves throughout this experiment. Copper sulfate stains the skin. All solutions will be placed into provided waste container upon completion.

Apparatus

The digital pH meter can be used as a voltmeter by pressing **mode** until the meter displays **mV** (rather than **pH** or **REL mV**).

Procedure

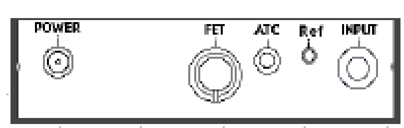
Part A: Oxidation-Reduction Reaction

1. Fill two 30-mL beakers about half full with 1 M copper sulfate solution.

2. Put a piece of silver foil in one beaker and a piece of zinc metal in the other. After a few minutes, record what you observe.

Part B: Measuring Voltages Produced by Voltaic Cells

1. Connect the leads to the Input Connector on the rear panel. If you hold the red and the black leads together, the meter should read ‘0.0 mV’.



2. To see if the meter is operating correctly, measure the voltage of a 1.5V battery.

3. Assemble copper and zinc half-cells. Fill one 30-mL beaker about 2/3 full with the blue 0.1 M copper sulfate solution. Fold a piece of the Cu foil over the beaker so that the foil is immersed in the solution. Make the zinc half-cell using 0.1 M zinc sulfate solution and a strip of Zn metal. Note: Oxide can be removed by “scratching” the metal with the leads.

4. Attach the red lead to the Cu foil and the black lead to the Zn metal. Make a salt bridge by soaking a strip of filter paper in KNO_3 solution. Join the half cells with the salt bridge.

5. Record the measured voltage. The stable icon shown will appear when the voltage is stable. The reading may be recorded at this time. Compare with the calculated voltage for a Cu vs. Zn cell.

Note: In all the voltaic cells in this experiment, copper will be the cathode.

6. Repeat for a Cu vs. Sn cell. You can use the same copper half-cell and the same salt bridge. Make the tin (Sn) half-cell using 0.1 M tin sulfate solution and a strip of Sn metal. Attach the red lead to the Cu foil and the black lead to the Sn metal. Join the half cells with the salt bridge. Resoak the salt bridge paper in KNO_3 solution if needed. Record the measured voltage, and compare with the calculated voltage for a Cu vs. Sn cell.

Part C: Electrolytic Cell and Copper Plating

1. Fill a 30-mL beaker about half full with the 1 M copper sulfate solution.

2. Connect one lead from the cathode (object you want to plate) to negative pole of the 9V battery. The other lead goes from the anode to the positive one. For the anode you can use a strip of copper metal or a penny dated before 1982. The cathode can be quarters, dimes, paper clips, anything metallic. Immerse electrodes in the solution and observe what happens. Note: Nickels are difficult to plate. If you remove the cathode from the solution you should see some copper deposited almost immediately. If not, check your connections.

3. Sketch and label the apparatus in the data/result section.

Data and Results (Electrochemistry)

*Name*____ *Date* ___________

Part A: Oxidation-Reduction Reaction

Reaction	Description of Reaction Mixture
Zn + Cu^{2+} (aq)	
Ag + Cu^{2+} (aq)	

Part B: Measuring Cell Potentials

Cell	Voltage Measured (mV)	*Voltage Measured (V)	Cell Potential Calculated (V)
Cu vs. Zn			
Cu vs. Sn			

* 1 V = 1000 mV

Part C Electroplating

Sketch the setup used for electroplating, labeling the DC source, cathode, anode and electrolyte solution.

Questions

1. Some measured voltages would be significantly different from the calculated values? Why would a Cu vs. Al measured voltage that you might perform in a lab experiment be different from the calculated value? (Hint: See Step B.3)

2. What would be needed to silver plate an object?

____

Instructor's Guide
(#25 *Electrochemistry*)

Data and Results (Electrochemistry)

Part A: Oxidation-Reduction Reaction

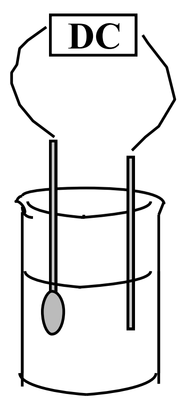
Reaction	Description of Reaction Mixture
Zn + Cu^{2+} (aq)	*zinc dissolves and copper is precipitated from solution*
Ag + Cu^{2+} (aq)	*no effect*

Part B: Measuring Cell Potentials

Cell	Voltage Measured (mV)	*Voltage Measured (V)	Cell Potential Calculated (V)
Cu vs. Zn	*1080*	*1.08*	1.10
Cu vs. Sn	*488*	*0.49*	0.48

* 1 V = 1000 mV

Part C: Electroplating



Answers to Questions

1. The cell potential for the aluminum half cell is listed in textbooks as –1.66 V for Al^{3+}/Al. So combined with Cu (0.34 V) the overall cell potential should be calculated to be around 2 V. However, the measured value that you would get from a lab experiment would not be close to that due to fact that aluminum is easily coated with oxides (therefore not pure). The Al to which textbooks refer is freshly deposited aluminum metal under nitrogen.

2. For the anode, you would need a strip of silver metal.

Instructor's Guide
(*Electrochemistry*)

Equipment and Materials (per group)

Items	Number/ amount	Comment
pH-voltmeters and power cords	1	
30-mL beaker	2	
Copper foil strip	1	Cu is copper
Zinc metal strip	1	Zn is zinc
Tin metal strip	1	Sn is tin
Silver foil strip	1	Ag is silver
Test leads	1 set	1 red and 1 black
1.5V battery	1	
9V battery	1	
Alligator clips	2	
0.1 M Zinc sulfate (aq)	100 mL	$ZnSO_4 \cdot 7H_2O$
0.1 M Copper sulfate (aq)	100 mL	$CuSO_4 \cdot 5H_2O$
0.1 M Tin sulfate (aq)	100 mL	$SnSO_4$
1 M Copper sulfate (aq)	100 mL	$CuSO_4 \cdot 5H_2O$
1 M Potassium nitrate (aq)	20 mL	KNO_3
Coarse filter paper strips	2	
Safety glasses	1	1 /student
Rubber gloves		1 box per class

Ideas/ Information
Solutions can be prepared by adding grams (*g*) of the compound to a volumetric flask and then diluting with distilled water to the mark on the flask.

Compound	Molarity (mol/L)	*g* compound for 1 L solution	*g* compound for 500 mL solution
zinc sulfate*	0.1	28.7	14.35
copper sulfate*	0.1	25	12.5
tin sulfate	0.1	21.5	10.75
copper sulfate*	1.0	249.6	124.8
potassium nitrate	1.0	101	50.5

* zinc sulfate is the heptahydrate: $ZnSO_4 \cdot 7H_2O$
* copper sulfate is the pentahydrate: $CuSO_4 \cdot 5H_2O$