Title: Electrochemistry

Target Audience: AP Chemistry students (possibly Honors)

Time: Three class periods (90 minutes each)

Prior Knowledge: Students should have covered chemical reactions including balancing, nomenclature, ionic and covalent compounds, acids and bases, and thermochemistry before beginning this unit.

Goals: At the end of the unit, students will:

- be able to define terminology such as oxidation, reduction, oxidation number, oxidizing agent, reducing agent, anode, cathode, voltaic cell, half-reaction, standard reduction potential
- understand the rules in determining oxidation numbers and apply them to redox reactions
- use the rules for balancing redox reactions and apply them to problems
- understand what a voltaic cell is and what information can be obtained from it
- be able to use a standard reduction table to calculate standard potentials for reactions
- perform redox reactions in the lab and calculate standard reduction potentials

Standards: PA Department of Education Standards for Science and Technology

3.2.10 Inquiry and Design

3.2.10A- Apply knowledge and understanding about the nature of scientific and technological knowledge

3.2.10B- Apply process knowledge and organize scientific and technological phenomena in varied ways

- describe materials using precise quantitative and qualitative skills based on observation
- Develop appropriate scientific experiments

3.2.10C- Apply the elements of scientific inquiry to solve problems

- conduct a multi step experiment
- organize experimental information using a variety of analytical methods
- judge the significance of experimental information in answering the question

3.2.10D- Identify and apply the technological design process to solve problems

3.4.10 Physical Science, Chemistry and Physics

3.4.10A- Explain concepts about the structure and properties of matter

- know that atoms are composed of even smaller subatomic structures whose properties are measurable
- describe various types of chemical reactions to applying the laws of conservation of mass and energy

3.4.10B- Analyze energy sources and transfers of heat

- use knowledge of chemical reactions to generate an electrical current
- determine the efficiency of chemical systems by applying mathematical formulas
- evaluate energy changes in chemical reactions
- use knowledge of conservation of energy and momentum to explain common phenomena

3.7.10- Technological Devices

3.7.10A- Identify and safely use a variety of tools, basic machines, materials and techniques to solve problems and answer questions

3.7.10B- Apply appropriate instruments and apparatus to examine a variety of objects and processes

- describe and use appropriate instruments to gather and analyze data
- apply accurate measurement knowledge to solve everyday problems

Day 1: Introduction to electrochemistry

Objectives – Students will be able to differentiate between oxidation and reduction. Students will be able to balance redox reactions in acidic solution.

Assessments – informal observation by walking around the classroom, collection and grading of homework assignment.

1. **Warm-up activity: Demonstration**- Reaction between copper (in the form of a penny) and concentrated nitric acid:

$$Cu(s) + 4HNO_3 \rightarrow Cu(NO_3)_2(aq) + 2NO_2(g) + 2H_2O(l)$$

This reaction should get the students' attention and introduced to redox reactions. Have students write observations in their notebook while the reaction is progressing. Put the equation on the board to illustrate what is happening. During the reaction, a brown gas is produced and the solution changes to a blue color as copper is oxidized to copper (II) nitrate. **Make sure to use the proper safety precautions since you are using concentrated acid.** Tell the students that the demo was an example of a redox reaction and that we will be starting that chapter today.

2. **Student-Guided Activity: POGIL**: Redox Reactions (Appendix 1)

Have students get into groups of three or four. Hand out the POGIL "Redox Reactions" and have students work cooperatively to complete it. Walk around the room and monitor student progress. When all students finish, review answers together and correct any misconceptions.

3. **Teacher-guided activity**- Balancing Redox reactions

Teacher will hand out "Rules for balancing redox reactions" (Appendix 2). Have students follow along as the teacher balances the following reaction on the promethean board. Go through step-by-step and ask the class if there are any questions as each step is performed.

$$MnO_4^- (aq) + C_2O_4^{2-} (aq) \rightarrow Mn^{2+} (aq) + CO_2 (aq)$$

4. **Homework**- Balancing redox reactions

If time allows, students may begin to work on "Balancing redox reactions" worksheet (Appendix 3). Students should finish for homework.

Day 2- Review balancing redox equations and galvanic cells

Objectives – Students will be able to balance redox reactions. Students will understand what a voltaic cell is and what it is used for. Students will be able to calculate standard reduction potentials with the use of the standard reduction table.

Assessments- informal observation by walking around the room, student completion of worksheet

1. **Warm-up activity**- The following redox equation will be put on the chalkboard:

$$Cu\ (s) + NO_3^-\ (aq) \rightarrow Cu^{2+}\ (aq) + NO_2\ (g)$$

Students will balance the reaction in their notebook individually. Walk around the room and monitor their progress. When students are finished, have a volunteer come to the board and go through the steps of balancing. Ask if there are any questions. If students need additional help, continue with more redox reactions.

2. **Teacher lecture**- powerpoint presentation on galvanic cells (Appendix 4)

Students will take notes as the teacher presents galvanic cells and standard reduction potentials through the use of a computer and digital projector.

3. **Student group work**- Calculating Reduction Potentials
Hand out "Standard Reduction Potentials" (Appendix 5). Students will work in groups of four to complete "Electrochemical cells" worksheet (Appendix 6). When students are finished, volunteers will put answers on the board.

4. **Homework**- Pre-lab for tomorrow's lab- " Electrochemical Cells" (Appendix 7)
Students will read the lab and prepare the required components in their lab notebook (Title, Purpose, Background Information, Materials, Procedure, Data Table)

Day 3- Laboratory Experiment: Electrochemical Cells

Objective- Students will use appropriate laboratory techniques to create electrochemical cells and calculate standard reduction potentials.

Assessment- informal observation of lab experiment, collection of lab notebook for grading

1. **Warm-up activity**- Students will respond to the following question in their notebook.
While students are working, walk around and check the pre-labs.

What is an electrochemical cell? Name all of the parts involved and give a description of what each one does.

2. **Student Exploration**- Laboratory Experiment- “Electrochemical Cells” (Appendix 7)

Teacher led discussion of laboratory procedures. Go through lab and show students how to set up apparatus. Ask students if there are any questions before they begin. Students will complete experiment with their lab partner. They should follow all safety precautions and experimental procedures. When they finish, equipment should be cleaned up and put away. Lab stations need to be wiped down. Students may begin to work on conclusion questions.

3. **Homework-** Finish lab questions

Appendix 1

Redox Reactions

Oxidation-reduction reactions, otherwise known as "redox" reactions, involve the transfer of electrons. These two processes are always paired together.

Part I- Oxidation numbers

Model 1: Basic information

Oxidation= loss of electrons (e^-)
Reduction= gain of electrons (e^-)

Half-reactions

(1) $Na \rightarrow Na^+ + e^-$
(2) $Mg^{2+} + 2e^- \rightarrow Mg$

Memory device:
OIL RIG
Oxidation **I**s **L**oss
Reduction **I**s **G**ain

Critical Thinking Questions:

1.) Which equation (1 or 2) represents an oxidation reaction? (Hint: Which is losing negative charge, Na or Mg^{2+}?)

2.) Which equation (1 or 2) represents a reduction reaction? (Hint: Which is gaining negative charge, Na or Mg^{2+}?)

3.) Are the electrons that appear on the products side of Equation 1 gained or lost by the sodium?

4.) Are the electrons that appear on the reactants side of Equation 2 gained or lost by the magnesium ion?

To combine Equations 1 and 2 into one equation, there needs to be the same number of electrons involved in both processes.

5.) a.) What number should equation 1 be multiplied by to match the number of electrons in equation 2?

b.) Rewrite Equation 1 after the multiplication takes place. Remember to multiply the number through the entire equation. Rewrite equation 2 below equation 1. Add both equations together to get the net equation.

(1) _____ → _____

(2) _____ → _____

(1+2) _____ → _____

c.) Do the electrons cancel out completely when adding (1) and (2)? If not, go back to part a.

d.) Will the reaction in part b (the sum) occur based on the activity series of metals?

Part II: Molecular compounds

It is easy to tell what is getting oxidized and reduced when looking at an ionic equation (Model 1). For molecular equations, oxidation numbers will need to be assigned to tell what is getting oxidized and reduced.

Model 2: Molecular compounds

An oxidation number (or oxidation state) is a hypothetical charge assigned to an atom. For ions, the oxidation number equals the charge of the ion.
Example 1: NaCl the oxidation number for Na^+ =+1 and Cl^- = -1
Example 2: H_2O the oxidation number for H = +1 and O = +2

Critical Thinking Questions

6.) Keeping in mind that the oxidation numbers add up to zero for a compound (and add up to the overall charge for a polyatomic ion), assign oxidation numbers to each element in the following:

a.) CO_2 (Hint: if oxygens were ions, that charge would each one be?)

b.) SO_4^{-2}

c.) MnO_4^-

7.) Two substances in the following reaction represent elements that exist as atoms. Write a zero above each of those elements, showing their oxidation number is zero.

$Mg\ (s) + 2HCl\ (aq) \rightarrow MgCl_2\ (aq) + H_2\ (g)$

Assign oxidation numbers to the rest of the elements. Remember to assign them above each element in the equation. If more than one atom/ion occurs just write the oxidation number for one atom/ion above the symbol. (Example: Write +1 for H in HCl, even though there are two of them in the balanced equation).

a.) How does magnesium change in oxidation number in the reaction above? Does its oxidation number increase or decrease? Is it being oxidized or reduced?

b.) How does hydrogen change in oxidation number in the reaction above? Does its oxidation number increase or decrease? Is it being oxidized or reduced?

c.) How does chlorine change in oxidation number in the reaction above? Does its oxidation number increase or decrease? Is it being oxidized or reduced?

Model 3:

In all redox reactions, one substance is oxidized and one is reduced. The substance that makes it possible for another substance to be oxidized is called the **oxidizing agent.** The oxidizing agent removes electrons from another substance by acquiring them itself; therefore, the oxidizing agent is reduced. Similarly, a **reducing agent** is a substance that gives up electrons, thereby causing another substance to be oxidized. The reducing agent is oxidized in the reaction.

8.) In question 7, indicate which substance is the oxidizing agent and which is the reducing agent.

Problems

1. The nickel-cadmium (nicad) battery, a rechargeable "dry cell" used in battery-operated devices, uses the following redox reaction to generate electricity:

$Cd\ (s) + NiO_2\ (s) + 2H_2O \rightarrow Cd\ (OH)_2\ (s) + Ni(OH)_2(s)$

Identify the substances that are oxidized and reduced, and indicate which are oxidizing agents and which are reducing agents.

2.) Identify the oxidizing and reducing agents in the following reaction:

$2H_2O$ (l) + Al (s) + MnO_4^- → $Al(OH)_4^-$ (aq) + MnO_2(s)

Appendix 2

Balancing Equations by the Method of Half-reactions

The use of half-reactions provides a general method for balancing oxidation-reduction equations. For balancing a redox reaction that occurs in acidic aqueous solution, use the following rules:

1.) Assign oxidation states to see which atoms are gaining electrons and which are losing electrons.
2.) Divide the equation into two half-reactions, one for oxidation and the other for reduction.
3.) Balance each half-reaction:
    a. First, balance the elements other than H and O.
    b. Next, balance the O atoms by adding H_2O as needed.
    c. Then, balance the H atoms by adding H^+ as needed.
    d. Finally, balance the charge by adding e^- as needed.
4.) Multiply the half-reactions by integers if necessary so that the number of electrons lost in one half-reaction equals the number gained in the other.
5.) Add the two half-reactions and wherever possible, simplify by canceling species appearing on both sides of the combined equation.
6.) Check to make sure that the number of atoms of each kind on the left side of equation is the same as the number on the right side.
7.) Check to make sure that the total charge on the left is the same as the total charge on the right.

Appendix 3

Name _____ Balancing redox reactions

Complete and balance the following redox reactions. Assume they take place in an acidic environment.

1.) $Cr_2O_7^{2-}$ (aq) + I^- (aq) → Cr^{3+} (aq) + IO_3^- (aq)

2.) As_2O_3 (s) + NO_3^- (aq) → H_3AsO_4 (aq) + N_2O_3 (aq)

Appendix 4- See attached powerpoint

Appendix 5- See attached Reduction potential table

Appendix 6

Electrochemical Cells & Potentials **Name_____**

1. What is an electrochemical cell?

2. List the components of a voltaic cell, and describe the function of each.

3. List the following metals in order from more readily oxidized to least readily oxidized: Hg,Cu, Ni, Mg, Al, Pb.

4. When a strip of copper is dipped into a solution of iron (II) sulfate will it be oxidized? Why or why not?

5. What is the electrical potential of a voltaic cell?

6. Write each half-reaction for the following reaction. Then, find the standard cell potential for the reaction.

   a. $Cu(s) + 2H^{+}(aq) \rightarrow Cu^{2+}(aq) + H_2(g)$

   b. $2\ Ag(s) + Fe^{2+}(aq) \rightarrow 2\ Ag^{+}(aq) + Fe(s)$

7. The standard cell potential for a reaction is –0.74 volts. What does this mean?

8. Calculate the standard cell potential to determine if this redox reaction will occur spontaneously.
   $3\ Zn^{2+}(aq) + 2\ Cr(s) \rightarrow 3\ Zn(s) + 2\ Cr^{3+}(aq)$

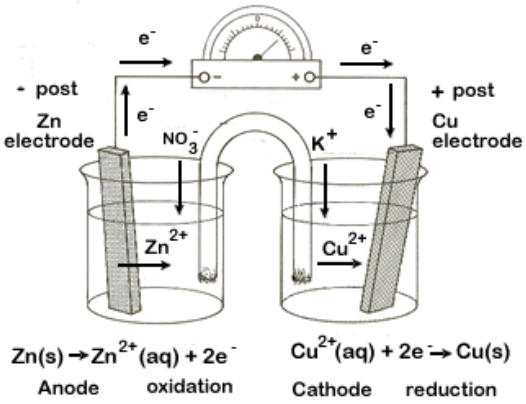
Electrochemistry- *http://www.saskschools.ca/curr_content/chem30_05/6_redox/labs/electrochem_cells.htm)*

Electrochemical Cells

OVERVIEW

In an electrochemical cell, chemical energy is converted into electrical energy. This is accomplished by using a spontaneous chemical reaction to generate an electric current, which we can simply define here as electrons traveling though a wire.

To create the electrochemical cell two half-reactions will be set up in different containers. In one, an oxidation reaction will be used to generate a source of electrons. These free electrons will travel, through an external circuit, to the second container and will cause the reduction reaction to occur. The final requirement for our complete electrochemical cell will be a salt bridge that will permit ions to flow between the two half-cells, thus maintaining electrically neutral solutions.



PURPOSE

- To build several electrochemical cells.

SAFETY

- Follow normal lab safety guidelines. There are no specific safety hazards for this lab.

EQUIPMENT AND MATERIALS

- 250-mL beakers, 2
- glass U-tube for the salt bridge
- cotton plugs for the salt bridge
- copper wires, 2, insulated, with alligator clips
- steel wool to clean electrodes
- DC voltmeter
- metal electrodes: copper, zinc, lead

Solutions

- 0.5M $Cu(NO_3)_2$
- 0.5M $Zn(NO_3)_2$
- 0.5M $Pb(NO_3)_3$
- 0.5M KNO_3 – for the salt bridge

PROCEDURE

1. Each half-cell will be created by placing a metal electrode in an electrolytic solution containing the same metal's ions. For example the copper electrode will be placed in a copper(II) nitrate solution.

   For each electrochemical cell you create you will require two half-cells. Set these cells beside each other – they will be connected by the U-tube.

2. Fill a beaker about two-thirds full of the electrolytic solution. Clean the electrode using the steel wool, then place the electrode in its appropriate solution.

3. Clip one end of each copper wire to the two electrodes using the alligator clips.

4. Fill the U-tube with KNO_3 and stopper both ends with the cotton plugs. Turn the U-tube upside down and place one end in each half-cell.

5. Touch the other end of the copper wires to the voltmeter terminals. If the indicator on the voltmeter deflects in the wrong direction, switch the wires on the terminals. Read the highest voltage reading obtained – you'll need to do this quickly after connecting the wires to the voltmeter.

6. Repeat the experiment for other combinations of half-cells.

RESULTS

Copy a data table similar to the one shown below into your lab notebook and use it to record your results.

Half-cells		Voltage
$Cu\|Cu^{2+}$	$Zn\|Zn^{2+}$	
$Cu\|Cu^{2+}$	$Pb\|Pb^{2+}$	
$Zn\|Zn^{2+}$	$Pb\|Pb^{2+}$	

CONCLUSIONS AND QUESTIONS

1. For each electrochemical cell you created:

    a. Write out the two half-reactions for each electrochemical cell you created.
    b. Identify each half-reaction as oxidation or reduction.
    c. Identify each half-reaction as the anode and cathode.
    d. Indicate the direction of the flow of electrons

2. Using a Table of Standard Reduction Potentials, calculate the theoretical voltage for each cell.

3. Compare the voltages you obtained with the theoretical voltage for each cell. What are some reasons that would account for any differences?