| | |
|---|---|
| **Early Booklet E.C.:** | **+ 2** |
| **Unit 13 Hwk. Pts.:** | **/ 32** |
| **Unit 13 Lab Pts.:** | **/ 32** |
| Late, Incomplete, No Work, No Units **Fees**? | Y / N |

# Unit 13 – Redox Reactions & Electrochemistry

Ch. 19 & 20 of your book.

## Learning Targets for Unit 13

1.1 I can describe the process of oxidation and reduction.
1.2 I can identify oxidizing and reducing agents.
1.3 I can determine the oxidation number of an element in a compound.
1.4 I can interpret redox reactions in terms of change in oxidation state and transfer of elections.
1.5 I can use changes in oxidation number to balance redox equations.
1.6 I can balance net ionic redox equations using the oxidation number method.
1.7 I can describe a way to obtain electrical energy from a redox reaction.
1.8 I can identify the parts of a voltaic cell and explain how each part operates.
1.9 I can calculate cell potentials and determine the spontaneity of redox reactions.
1.10 I can describe the structure, composition and operation of the typical carbon-zinc dry cell battery.
1.11 I can distinguish between primary and secondary batteries, and give two examples of each type.
1.12 I can explain the structure and operations of the hydrogen oxygen fuel cell.
1.13 I can describe the process of corrosion of iron and methods to prevent corrosion.
1.14 I can describe how it is possible to reverse a spontaneous redox reaction in an electrochemical cell.
1.15 I can compare the reactions involved in the electrolysis of molten sodium chloride with those in the electrolysis of brine.
1.16 I can discuss the importance of electrolysis in the smelting and purification of metals.

## Unit Vocabulary for Unit 13

| | | | |
|---|---|---|---|
| Oxidation-reduction reaction | Redox reaction | Oxidation | Reduction |
| Oxidizing agent | Reducing agent | Oxidation-number method | Species |
| Half-reaction | Salt bridge | Electrochemical cell | Voltaic cell |
| Half-cell | Anode | Cathode | Reduction potential |
| Standard hydrogen electrode | | | |

| Possible 13.1 Pts.: 8 |
| --- |
| Late, Incomplete, No work, No Units Fee: - 1 - 2 - 3 |
| **Final Score: / 8** |

## 13.1 Problems – Oxidation and Reduction
## Section 19.1 of your textbook.

1. Explain why not all oxidation reactions involve oxygen.

Identify the oxidizing and reducing reactant in each of these redox equations.

2. $N_2 + 3H_2 \rightarrow 2NH_3$

3. $2Na + I_2 \rightarrow 2\ NaI$

Identify each of these half reactions as either oxidation or reduction.

4. $Al \rightarrow Al^{3+} + 3e^-$

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

Determine the oxidation number of each element in these compounds or ions.

6. $NO_3^-$

7. $N_2O$

8. $NF_3$

| Possible 13.2 Pts.: 5 |
| --- |
| Late, Incomplete, No work, No Units Fee: - 1 - 2 |
| **Final Score: / 5** |

## 13.2 Problems – Balancing Redox Reactions
## Section 19.2 of your textbook.

Balance these redox equations. Use your ions list to determine ionic compound charges

1. _____ $PbS$ + _____ $O_2$ → _____ $PbO$ + _____ $SO_2$

2. _____ $NH_3$ + _____ $CuO$ → _____ $Cu$ + _____ $N_2$ + _____ $H_2O$

Balance these redox equations. You will have to rewrite 4 to account for water's presence.

3. _____ $PbO_{(s)}$ + _____ $NH_{3(g)}$ → _____ $N_{2(g)}$ + _____ $H_2O_{(l)}$ + _____ $Pb_{(s)}$

4. _____ $Br^-_{(aq)}$ + _____ $NO_3^-{}_{(aq)}$ → _____ $BrO^-_{(aq)}$ + _____ $NO_{(g)}$ in acid solution

5. _____ $H_2C_2O_4$ + _____ $HAsO_2$ → _____ $CO_2$ + _____ $As$ + _____ $O_2$

| Possible 13.3 Pts.: 8 |
| --- |
| Late, Incomplete, No work, No Units Fee: - 1 - 2 - 3 |
| **Final Score: / 8** |

## 13.3 Problems – Voltaic Cells
## Section 20.1 of your textbook.

1. What is the function of a salt bridge in a voltaic cell?

Using your Reduction Potentials Resource, write the standard cell notation for each cell in which each of the following half-cells is connected to the standard hydrogen electrode.

2. $Zn \mid Zn^{2+}$

3. $2\ Hg \mid Hg_2^{2+}$

4. $Cu \mid Cu^{2+}$

5. $Al \mid Al^{3+}$

Calculate the cell potential for the following voltaic cells.

6. $2Ag^{+}_{(aq)} + Pb_{(s)} \rightarrow Pb^{2+}_{(aq)} + 2Ag_{(s)}$

7. $Mn_{(s)} + Ni^{2+}_{(aq)} \rightarrow Mn^{2+}_{(aq)} + Ni$

8. $I_{2(aq)} + Sn_{(s)} \rightarrow 2I^{-}_{(aq)} + Sn^{2+}_{(aq)}$

| Possible 13.4 Pts.: 11 |
| --- |
| Late, Incomplete, No work, No Units Fee: - 1 - 2 - 3 - 4 |
| **Final Score: / 11** |

## 13.4 Problems – Batteries and Electrolysis
## Section 20.2 - 20.3 of your textbook.

Lead-Acid Battery. Answer questions 1 – 4 based on this.

1. The reaction of the lead-acid battery is shown. Assign oxidation numbers to all elements.

$$Pb + PbO_2 + H_2SO_4 \rightarrow 2\ PbSO_4 + 2\ H_2O.$$

2. What substance is reduced?

3. What substance is oxidized?

4. What substances are produced in each reaction? Write half reactions for each.

5. **Galvanization**. What is galvanization? How does galvanizing iron protect it from corrosion?

Half-reactions for a lead-acid storage battery are below:

$PbO_{2(s)} + SO_4^{2-}{}_{(aq)} + 4H_3O^+{}_{(aq)} + 2e^- \rightarrow PbSO_{4(s)} + 6H_2O_{(l)}$ $\quad E^0 = +1.685$ V

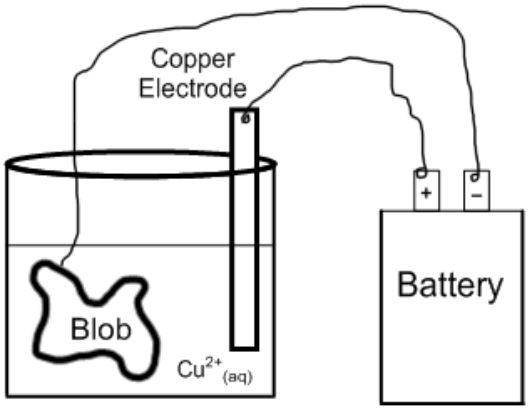
$PbSO_{4(s)} + 2e^- \rightarrow Pb_{(s)} + SO_4^{2-}{}_{(aq)}$ $\quad E^0 = -0.356$ V

6. What is the standard cell potential of one cell in a car battery?

7. Describe what happens at the anode in the electrolysis of $KI_{(l)}$.

8. Describe what happens at the cathode in the electrolysis of $KI_{(l)}$.

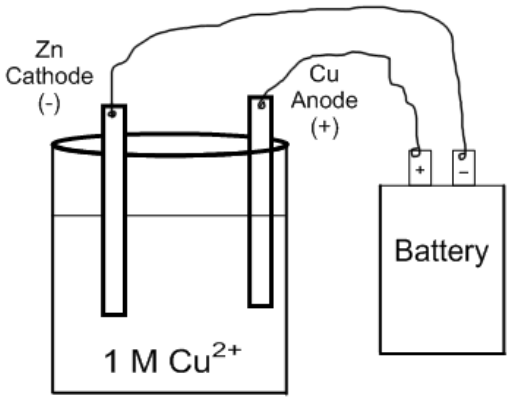
9. The figure shows a blob being electroplated with copper in an electrolytic cell. Where does oxidation occur? Explain your answer.




Answer the questions based on the picture below.

10. Which electrode grows? Write the reaction that occurs at this electrode.

11. Which electrode disappears? Write the reaction that occurs at this electrode.



| Chemistry | Lab 13.1: Pollution Lab | | | | | |
|---|---|---|---|---|---|---|
| Name: | | | | | | Correction Credit: Half |
| Lab Points: | E.C. | Missed: | Late, No Units, No Work Fee: | First Score: | Corrections: | Final Score: |
| 12 | 1 2 | | -1 -2 -3 | | | |

**Story**:

You are an investigator looking into the pollution of a sensitive waterway surrounded by three chemical plants. Each plant produces a particular chemical, and it is up to you to determine which chemical was released! Can you do it?

**Purpose:** Experimentally determine the identity of an unknown chemical pollutant based on observations, identify oxidized and reduced chemicals in reactions, and balance redox reactions.

**Theory:**

Oxidation-reduction (redox) reactions involve transferring electrons from one element to another one. As electrons are removed from an element, its oxidation number increases, and is oxidized. As the electrons are added to another element, that element's oxidation number diminishes and is reduced.

**Equipment**:

0.1 M silver nitrate ($AgNO_3$)
0.1 M Hydrochloric Acid (HCl)
0.1 M Zinc Sulfate ($ZnSO_4$)
Unknown Solution
Cu Wire
Fe Filings
Mg Ribbon
Pb Shot
Tweezers
2 - 12 Cell Reaction Plates

**Procedure**:

1. Place a piece of copper wire in four cells of the first column of your reaction plate.
2. Repeat Step 1 with the rest of the three metals, in columns 2, 3, and 4.
3. Count 8 drops of silver nitrate into each well of the first row.
4. Repeat Step 3 with the hydrochloric acid in the second row, the zinc sulfate in the third row, and the unknown solution in the fourth row.
5. Allow the reactions to proceed for 5 minutes, then describe your observations in the data table. Write NR for any cells that do not show evidence of a reaction.
6. Clean up your area by pouring all reaction plates and metals into the waste beaker in the front of the lab for further processing.

Data Table: (8 Points) Record your observations here: be thorough.

| | Cu Wire | Fe Filings | Mg Ribbon | Pb Shot |
|---|---|---|---|---|
| $AgNO_3$ (Silver Nitrate) | | | | |
| HCl (Hydrochloric Acid) | | | | |
| $ZnSO_4$ (Zinc Sulfate) | | | | |
| Unknown | | | | |

Questions/Calculations:

1. (2 Points) What is the identity of your unknown, and describe how you determined what it was.

2. (2 Points) Select two of the cells in which a chemical reaction occurred, and write a balanced chemical reaction for each. Label which element was oxidized and which was reduced – write down oxidation numbers for all elements to determine this. Assume that single replacement reactions happened, and that Cu $\rightarrow Cu^{2+}$, and Fe $\rightarrow Fe^{3+}$.

E.C. (2 Points): Balance the following redox reaction by whatever method you choose:

$Sb^{3+} + MnO_4^- \rightarrow SbO_4^{3-} + Mn^{2+}$ (Acidic Solution)

| Chemistry | Lab 13.2: Electrochemical Cell Lab | | | | |
|---|---|---|---|---|---|
| Name: | | | | | Correction Credit: Half |
| Lab Points: | Missed: | Late, No Units, No Work Fee: | First Score: | Corrections: | Final Score: |
| 14 | | -1 -2 -3 | | | |

Purpose:

Make an electrochemical cell and calculate electric potential.

Overview.

Electrochemical cells derive their potential from chemical reactions. Redox reactions between metals cause electrons to move through conductors in the cells, and during this process they are able to do work.

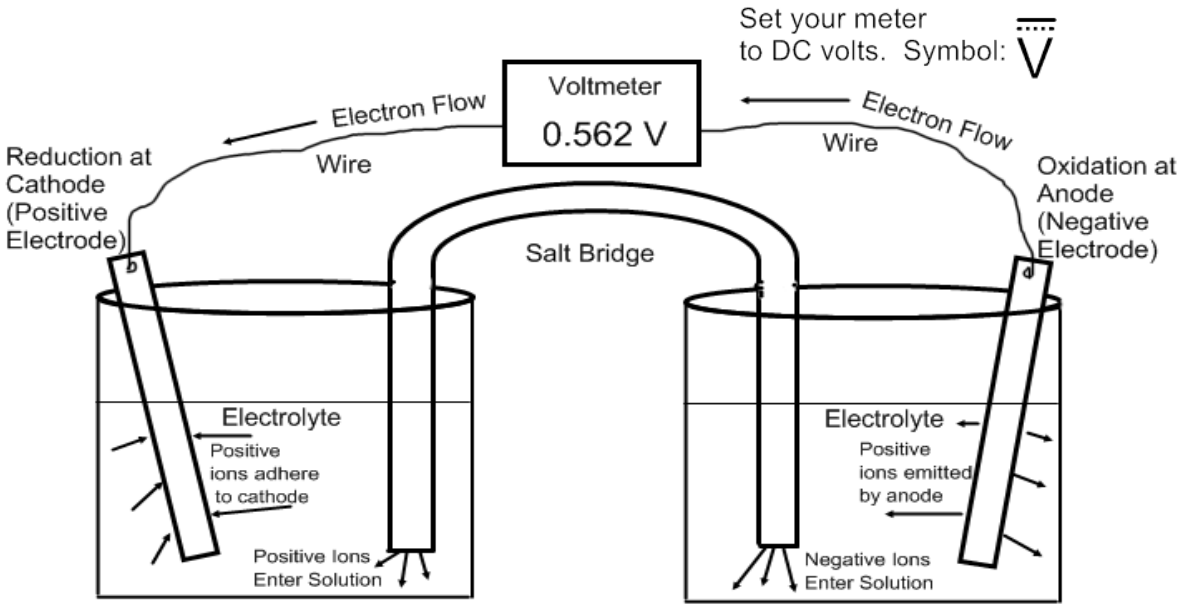
When the anode (negative electrode) is oxidized it changes from a neutrally charged solid to a positively charged ion, and goes into solution. This conversion to a positive ion causes it to lose electrons, which travel through conductors to the cathode (positive electrode). At the cathode, the opposite happens: positively charged metal ions in solution adhere to the cathode, absorbing electrons from it as they do so, thus obtaining a neutral charge and becoming solid.

To maintain a balance of charge during the movement of the electrons, the salt bridge allows ions from one solution to travel into the other one. If there were no balancing of ions, the electrochemical cell would not work for long.

In standard conditions, the total electric potential of a cell can be calculated using the following equation:

Equation 1: $E^0_{cell} = E^0_{reduction} - E^0_{oxidation}$

The resource section of your booklet has standard values for different metals.




Electrochemical Cell Diagram

**Materials:**

Plastic cups
Salt bridge
Voltmeter (Set to DC Volts!)
1.0 M NaCl

Pure Element Electrodes: Clean them first.

Zinc (Zn)
Copper (Cu)
Lead (Pb)
Iron (Fe)
Aluminum (Al)

Alloy Electrodes:

Brass (Cu/Zn)
Square Solder – (Sn/Pb/Zn)

Apparatus. Set up a basic cell in the following manner:

1. Fill two plastic cups halfway with 1.0 M NaCl electrolyte solution.
2. Place a CPVC salt bridge over the edges of the cups.
3. Choose any two of the electrodes and put one in each cup.
4. Touch the probes of the voltmeter to the electrodes. Make sure the meter is set to DC volts.
5. Check polarity – the voltage should read positive. If it doesn't, switch your probes. When the voltage reads positive, your cathode will be the POSITIVE (red) electrode, and your anode will be the NEGATIVE (black) electrode.
6. Clean up by pouring the solution back in the jug, and rinsing your electrodes off with water.

Electrical Potential (9 points).

Set up a cell, and test the potential of any two different metals. Make sure your voltage is positive, then record your data. Switch the metals out with different ones so that you will have made three unique cells by the time you're done. Report voltages to the hundredths place.

| Cell #: | Anode Metal (-) | Cathode Metal (+) | Voltage |
|---|---|---|---|
| 1. | | | |
| 2. | | | |
| 3. | | | |

Calculations

1. (1 Point) Which two metals produced the highest voltage and what was that voltage?

2. (2 Points) Using the resource section of your booklet, and Equation 1 from this lab, determine what the electrical potential for the highest cell containing PURE ELEMENTS (not alloys) would be. You can see which electrodes are pure by looking at the Materials list.
   Assume: $Fe^{3+} + 3e^- \rightarrow Fe$, $Cu^{2+} + 2\ e^- \rightarrow Cu$, and $Pb^{2+} + 2\ e^- \rightarrow Pb$.

3. (2 Points) Use standard cell notation (use your notes to guide you here) to describe your cell from Question 2. You will have to determine which electrode is reduced and which is oxidized to do this.

| Chemistry | Lab 13.3: Electroplating Lab | | | | |
|---|---|---|---|---|---|
| Name: | | | | | Correction Credit: Half |
| Lab Points: | Missed: | Late, No Units, No Work Fee: | First Score: | Corrections: | Final Score: |
| 6 | | -1 -2 | | | |

Purpose:

Use the process of electroplating to study and understand electrochemistry.
In this lab, all you have to do to get credit is to show me something you have electroplated.

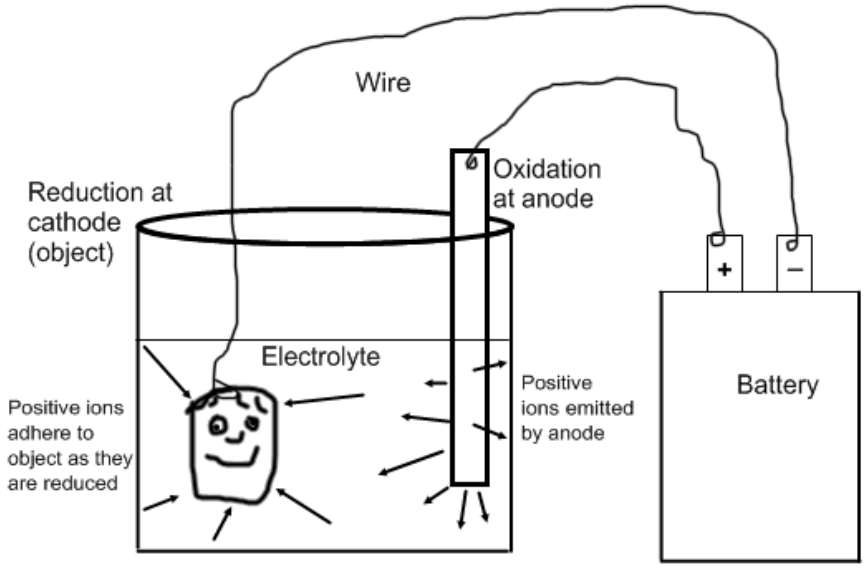
Overview.

Electroplating is a method commercially used to coat objects with a layer of other metal. For example, silver plating is used to coat steel, such that a silvery appearance is achieved without the effort or expense of making the object out of pure silver.

Reduction of metal ions by a charged cathode will lead to the cathode being coated. At the anode, oxidation causes the metal to ionize and go into solution. Eventually, the anode will deteriorate to the extent that there is none left.

**Materials:**

Plastic Cups
6 V Battery
2 Alligator Clips
Alcohol and Steel Wool
1.0 M $ZnSO_4$ Solution
1.0 M $CuSO_4$ Solution
Zinc (Zn) Electrode
Copper (Cu) Electrode
Object(s) to Electroplate



Electrolytic Cell Diagram

Procedure:

1. Clean your object with alcohol and paper towels. If there is a lot of grime or a protective coating on your object, you may have to use steel wool and wear it off.
2. You can plate your object(s) with zinc, copper, or both, but you will need two separate electrochemical cells.
3. Fill your cell with either the $ZnSO_4$ or $CuSO_4$ solutions.
4. Connect an electrode (Zn in $ZnSO_4$ or Cu in $CuSO_4$) to the positive terminal of the battery with an alligator clip, then put the electrode into the cell. DON'T touch the alligator clip to the electrolyte.
5. Connect your object, via alligator clip, to the negative terminal of the battery.
6. Keeping your object as far away from the electrode as possible, dip it in the solution and watch what happens – you may want to move the object a bit to apply a more even coating. Also, if you reposition the alligator clip you will get a better product. Having the object too close to the electrode can cause it to tarnish as other redox reactions occur.

Cleanup:

1. Leave your cells full of solution when you are done, and make sure to disconnect all wires from your battery.

2. Wipe off all residue that may have spilled during your experiment with the sponges.

Scoring:

1. Show me the thing you electroplated and I'll check you off.

| | |
|---|---|
| **Points Possible:** | **10** |
| **Late/Inc. Fee: -1 -2 - 3** | |
| **Final Score:** | **/ 10** |

## Unit 14 Review – Redox Reactions and Electrochemistry

This serves as test preparation for the Unit 10 Test. Points earned are based on completion, and we will go over any questions you have during the review.

# Insert Review Here!