Oakland Schools Chemistry Resource Unit

Electrochemistry

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<u>Electrochemistry</u>

Content Statement:

C5.6x Reduction/Oxidation Reactions

Chemical reactions are classified according to the fundamental molecular or submolecular changes that occur. Reactions that involve electron transfer are known as oxidation/reduction (or "redox").

Content Expectations:

C5.6a **-** Balance half-reactions and describe them as oxidations or reductions.

C5.6b **-** Predict single replacement reactions (omitted in this section).

C5.6c **-** Explain oxidation occurring when two different metals are in contact.

C5.6d **-** Calculate the voltage for spontaneous redox reactions from the standard reduction potentials.

C5.6e **-** Identify the reactions occurring at the anode and cathode in an electrochemical cell.

Instructional Background Information:

Redox:

Redox is an abbreviation that combines the concepts of "Reduction" and "Oxidation." Redox reactions result from the competition for electrons between atoms. This is indicated by a change in the expected oxidation numbers of atoms as they progress from reactants to products. Oxidation and reduction occur SIMULTANEOUSLY and cannot occur without the other [*conservation of mass and conservation of charge*]

Oxidation State:

Oxidation number is the charge an atom possesses, or appears to possess, when electrons are counted according to certain arbitrary rules

Electrons shared between 2 unlike atoms are counted as belonging to the more electronegative atom

Electrons shared between 2 like atoms are counted being shared equally between the sharing atoms

Summary of Rules for determining Oxidation State of Atoms

The oxidation number of a FREE element is zero i.e. Na^{0}

The oxidation state of monatomic ions is the charge of the ion i.e. Na^{+}

Oxygen's oxidation number is –2 (except in "peroxides")
i.e. $H^{+1}\mathbf{O}^{\mathbf{-2}}H^{+1}$ i.e. $H_2^{+1}O_2^{-1}$

Hydrogen's oxidation number is +1 (except in "hydrides")
i.e. $H^{\mathbf{+1}}O^{-2}H^{\mathbf{+1}}$ i.e. $Na^{+1}H^{\mathbf{-1}}$

Group IA (1) elements are always +1 [*alkali metals*]

Group IIA (2) elements are always +2 [*alkaline earth metals*]

Ions of Halogens (Group 17 or VIIA) in binary compounds = -1

The algebraic sum of the + and – oxidation states in a polyatomic ion equals the charge on the ion (see Reference Table F)

The algebraic sum of the + and – oxidation states in a compound = 0

Oxidation Numbers are based on the stability of atoms and compounds: Electrical neutrality and a complete valence.

Reduction:
Any chemical change in which there is a DECREASE in oxidation number of an atom. There is a GAIN of electrons from an atom in the reaction.

1. The particle or atom that increases in oxidation number is said to be Reduced
2. The particle or atom that is oxidized acts as a oxidizing agent in a redox reaction

$$2\ Na^{0} + \mathbf{S^{0}} \longrightarrow Na^{+1}\mathbf{S^{-2}}$$

- The element **S** gains two electrons
- The oxidation number of **S** decreases from 0 to -2
- **S** is reduced in the redox reaction
- **S** is the oxidizing agent in the redox reaction

Oxidation:

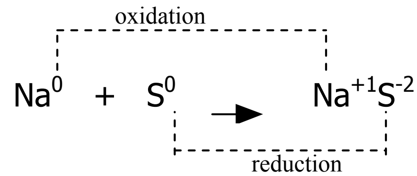
Any chemical change in which there is an INCREASE in oxidation number of an atom. There is a LOSS of electrons from an atom in the reaction.

1. The particle or atom that increases in oxidation number is said to be oxidized
2. The particle or atom that is oxidized acts as a reducing agent in a Redox reaction

$$2\ Na^{0} + S^{0} \longrightarrow Na^{+1}S^{-2}$$

- The element Na loses an electron
- The oxidation number of Na increases from 0 to +1
- Na is oxidized in the redox reaction
- Na is the reducing agent in the redox reaction

Half Reactions:
REDOX reactions are considered in 2 parts and each reaction of the 2 parts is called a ½ reaction (half rxn)

1. 

Reduction half reaction $S^0 + 2\ e^- \rightarrow S^{-2}$

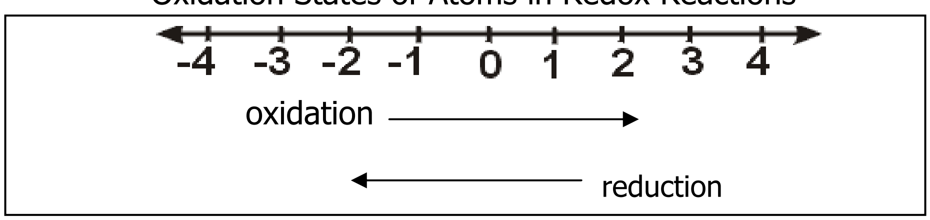
- This half reaction gains 2e-

Oxidation half reaction $2(\ Na^0 \rightarrow Na^{+1} + e^-\ ) = 2\ Na^0 \rightarrow 2\ Na^{+1} + 2\ e^-$

- This half reaction (oxidation of sodium) loses 1 e- by itself
- Since the number of electrons transferred between the two half reactions must be balanced to conserve charge, the number of moles of sodium being oxidized must be doubled to lose 2 e- (matching the 2 e- gained in the reduction half reaction)

- Overall, the reduction half reaction loses 2e-, therefore, the oxidation half reaction must gain 2e-

Oxidation States of Atoms in Redox Reactions



Balancing Redox Reactions: (Electron transfer method/Oxidation state method)

1. Write formulas for products and reactants in "skeletal" (unbalanced) equations.

$$HNO_{3\,(aq)} + H_3PO_{3\,(aq)} \rightarrow NO_{(aq)} + H_3PO_{4(aq)} + H_2O_{(l)}$$

2. Assign oxidation numbers (states) to each element on both sides of the equation.

$$H^{+1}\mathbf{N^{+5}}O_3^{-2}{}_{(aq)} + H_3^{+1}\mathbf{P^{+3}}O_3^{-2}{}_{(aq)} \rightarrow \mathbf{N^{+2}}O^{-2}{}_{(aq)} + H_3^{+1}\mathbf{P^{+5}}O_4^{-2}{}_{(aq)} + H_2^{+1}O^{-2}{}_{(l)}$$

- Notice that Nitrogen and Phosphorus change oxidation states from reactants to products. This reveals the redox reaction.

3. Select the atom that is oxidized (loses e-, increases oxidation number from reactant to products) and select the atom that is reduced (gains e-, decreases in oxidation number from reactants to products)

<u>Oxidation half reaction</u> $P^{+3} \rightarrow P^{+5} + 2\ e^-$

<u>Reduction half reaction</u> $N^{+5} + 3\ e^- \rightarrow N^{+2}$

4. Write balanced half-reactions (oxidation and reduction)
5. Inspect the half-reactions, checking for conservation of mass (same number of atoms on both sides) and conservation of charge (same total charge on both sides)
6. Combine the half-reactions so that the number of electrons lost in oxidation equals the number of electrons gained in reduction

<u>Oxidation half reaction</u> $\mathbf{3}\ (P^{+3} \rightarrow P^{+5} + 2\ e^-)$ … 6 e- lost

<u>Reduction half reaction</u> $\mathbf{2}\ (N^{+5} + 3\ e^- \rightarrow N^{+2})$ … 6 e- gained

7. Return to the overall chemical equation and add in appropriate coefficients (based on the half-reactions) to atoms that were oxidized and reduced.

$$\mathbf{2}\ HNO_{3\,(aq)} + \mathbf{3}\ H_3PO_{3\,(aq)} \rightarrow \mathbf{2}\ NO_{(aq)} + \mathbf{3}\ H_3PO_{4(aq)} + H_2O_{(l)}$$

8. Finish inspecting the entire chemical equation ensuring conservation of mass and charge for all atoms (reactants and products).

11 hydrogen atoms on each side of the equation
2 nitrogen atoms on each side of the equation
3 phosphorus atoms on each side of the equation
15 oxygen atoms on each side of the equation

Example: Give students the name of the elements and compounds only and have them determine half reactions and balance the overall equation:

Copper (s) + Nitric Acid → Copper II Nitrate (aq) + Nitrogen II Oxide + Water

$$3\ Cu\ (s) + 8\ HNO_3\ (aq) \rightarrow 3\ Cu(NO_3)_2\ (aq) + 2\ NO + 4\ H_2O\ (s)$$

Oxidation ½ reaction: $3\ (\ Cu\ (s) \rightarrow Cu^{2+}\ (aq) + 2\ e^-\ )$

Reduction ½ reaction: $2\ (\ N^{+5}\ (aq) + 3\ e^- \rightarrow N^{+2}\ (s)\ )$

Electrochemistry:

Electrochemistry translates the chemical energy of a reduction–oxidation reaction into electrical energy.

1. When the substances involved in oxidation and reduction half–reactions are physically separated, it is called an electrochemical cell.
2. Each half reaction occurs on the surface of an electrically conductive solid called an electrode.
3. Each electrode is immersed in a solution containing ions needed for the half–reaction.
4. The electrodes are connected by a wire so that electrons can move from the oxidation half–reaction to the reduction half–reaction.
5. The solutions are connected by a salt bridge so that ions can move between solutions.
6. In an electrochemical cell the chemical potential energy can be harnessed as the substances undergoing oxidation push electrons through the wire to the substances undergoing reduction.

Anode:
The anode is the electrode in the electrochemical cell where oxidation occurs. Anions are attracted to the anode because the anode is positively charged since it (the anode) is losing electrons in the half reaction. i.e. $Mg^0 \rightarrow Mg^{+2} + 2\ e-$

Cathode:
The cathode is the electrode in the electrochemical cell where reduction occurs. Cations are attracted to the cathode because the cathode is negatively charged since it (the cathode) is gaining electrons in the half reaction. i.e. $Cu^{+2} + 2\ e- \rightarrow Cu^0$
[*notice that the oxidation state of copper is "reduced"*]

Salt Bridge:
The salt bridge moves ions into those solutions to maintain electric neutrality without mixing of solutions.

Standard Reduction Potential (SRP)
By tradition, SRP tables such as the one shown on the next page list potentials for reduction half–reactions at standard state (1 *M*, 25 °C and 10^5 Pa).

1. Unfortunately, since one–half of a reaction does not occur without the other half, it is impossible to measure electrode potentials directly. The standard cell potential is the difference between the reduction potential of the cathode and the reduction potential of the anode.
2. The term "standard" comes from using the standard hydrogen electrode as the oxidation half–reaction: [$2\ H^+ + 2\ e^- \rightleftharpoons H_2$]. This electrode is defined as having an electrode potential of exactly 0 V.
3. To obtain standard oxidation potentials, the opposite sign is used.
4. These tabulated values can be used to determine the standard cell potential for any electrochemical cell.

Reduction potential decreases as one goes DOWN this table. Notice that non-metals make the strongest oxidizing agents.

strong oxidizing agents ↑

Standard Reduction Potentials at 298K, 1M, 1atm

HALF-REACTION	E° (V)
$F_{2(g)} + 2\,e^- \rightarrow 2\,F^-_{(aq)}$	+2.87
$O_{3(g)} + 2\,H^+_{(aq)} + 2\,e^- \rightarrow O_{2(g)} + H_2O_{(l)}$	+2.07
$Co^{3+}_{(aq)} + e^- \rightarrow Co^{2+}_{(aq)}$	+1.82
$H_2O_{2(aq)} + 2\,H^+_{(aq)} + 2\,e^- \rightarrow 2\,H_2O_{(l)}$	+1.77
$PbO_{2(s)} + 4\,H^+_{(aq)} + SO_4^{2-}{}_{(aq)} + 2\,e^- \rightarrow PbSO_{4(s)} + 2\,H_2O_{(l)}$	+1.70
$Ce^{4+}_{(aq)} + e^- \rightarrow Ce^{3+}_{(aq)}$	+1.61
$MnO_4^-{}_{(aq)} + 8\,H^+_{(aq)} + 5\,e^- \rightarrow Mn^{2+}_{(aq)} + 4\,H_2O_{(l)}$	+1.51
$Au^{3+}_{(aq)} + 3\,e^- \rightarrow Au_{(s)}$	+1.50
$Cl_{2(g)} + 2\,e^- \rightarrow 2\,Cl^-_{(aq)}$	+1.36
$Cr_2O_7^{2-}{}_{(aq)} + 14\,H^+_{(aq)} + 6\,e^- \rightarrow 2\,Cr^{3+}_{(aq)} + 7\,H_2O_{(l)}$	+1.33
$MnO_{2(s)} + 4\,H^+_{(aq)} + 2\,e^- \rightarrow Mn^{2+}_{(aq)} + 2\,H_2O_{(l)}$	+1.23
$O_{2(g)} + 4\,H^+_{(aq)} + 4\,e^- \rightarrow 2\,H_2O_{(l)}$	+1.23
$Br_{2(l)} + 2\,e^- \rightarrow 2\,Br^-_{(aq)}$	+1.07
$NO_3^-{}_{(aq)} + 4\,H^+_{(aq)} + 3\,e^- \rightarrow NO_{(g)} + 2\,H_2O_{(l)}$	+0.96
$2\,Hg^{2+}_{(aq)} + 2\,e^- \rightarrow Hg_2^{2+}{}_{(aq)}$	+0.92
$Hg_2^{2+} + 2\,e^- \rightarrow 2\,Hg_{(l)}$	+0.85
$Ag^+_{(aq)} + e^- \rightarrow Ag_{(s)}$	+0.80
$Fe^{3+}_{(aq)} + e^- \rightarrow Fe^{2+}_{(aq)}$	+0.77
$O_{2(g)} + 2\,H^+_{(aq)}\ 2\,e^- \rightarrow H_2O_{2(aq)}$	+0.68
$MnO_4^-{}_{(aq)} + 2\,H_2O_{(l)} + 3\,e^- \rightarrow MnO_{2(s)} + 4\,OH^-_{(aq)}$	+0.59
$I_{2(s)} + 2\,e^- \rightarrow 2\,I^-_{(aq)}$	+0.53
$O_{2(g)} + 2\,H_2O + 4\,e^- \rightarrow 4\,OH^-_{(aq)}$	+0.40
$Cu^{2+}_{(aq)} + 2\,e^- \rightarrow Cu_{(s)}$	+0.34
$AgCl_{(s)} + e^- \rightarrow Ag_{(s)} + Cl^-_{(aq)}$	+0.22
$SO_4^{2-}{}_{(aq)} + 4\,H^+_{(aq)} + 2\,e^- \rightarrow SO_{2(g)} + 2\,H_2O_{(l)}$	+0.20
$Cu^{2+}_{(aq)} + e^- \rightarrow Cu^+_{(aq)}$	+0.15
$Sn^{4+}_{(aq)} + 2\,e^- \rightarrow Sn^{2+}_{(aq)}$	+0.13
$2\,H^+_{(aq)} + 2\,e^- \rightarrow H_{2(g)}$	0.00
$Pb^{2+}_{(aq)} + 2\,e^- \rightarrow Pb_{(s)}$	-0.13
$Sn^{2+}_{(aq)} + 2\,e^- \rightarrow Sn_{(s)}$	-0.14
$Ni^{2+}_{(aq)} + 2\,e^- \rightarrow Ni_{(s)}$	-0.25
$Co^{2+}_{(aq)} + 2\,e^- \rightarrow Co_{(s)}$	-0.28
$PbSO_{4(s)} + 2\,e^- \rightarrow Pb_{(s)} + SO_4^{2-}{}_{(aq)}$	-0.31
$Cd^{2+}_{(aq)} + 2\,e^- \rightarrow Cd_{(s)}$	-0.40
$Fe^{2+}_{(aq)} + 2\,e^- \rightarrow Fe_{(s)}$	-0.44
$Cr^{3+}_{(aq)} + 3\,e^- \rightarrow Cr_{(s)}$	-0.74
$Zn^{2+}_{(aq)} + 2\,e^- \rightarrow Zn_{(s)}$	-0.76
$2\,H_2O_{(l)} + 2\,e^- \rightarrow H_{2(g)} + 2\,OH^-_{(aq)}$	-0.83
$Mn^{2+}_{(aq)} + 2\,e^- \rightarrow Mn_{(s)}$	-1.18
$Al^{3+}_{(aq)} + 3\,e^- \rightarrow Al_{(s)}$	-1.66
$Be^{2+}_{(aq)} + 2\,e^- \rightarrow Be_{(s)}$	-1.85
$Mg^{2+}_{(aq)} + 2\,e^- \rightarrow Mg_{(s)}$	-2.37
$Na^+_{(aq)} + e^- \rightarrow Na_{(s)}$	-2.71
$Ca^{2+}_{(aq)} + 2\,e^- \rightarrow Ca_{(s)}$	-2.87
$Sr^{2+}_{(aq)} + 2\,e^- \rightarrow Sr_{(s)}$	-2.89
$Ba^{2+}_{(aq)} + 2\,e^- \rightarrow Ba_{(s)}$	-2.90
$K^+_{(aq)} + e^- \rightarrow K_{(s)}$	-2.93
$Li^+_{(aq)} + e^- \rightarrow Li_{(s)}$	-3.05

strong reducing agents ↓

Oxidation potential decreases as one goes UP this table. Notice that metals make the strongest reducing agents.

Calculating Potentials in Half Reactions:

1. Consider the example below:

$$Mg\ (s) + Cu^{2+}\ (aq) \rightarrow Mg^{2+}\ (aq) + Cu\ (s)$$

- In this reaction magnesium (Mg) is being oxidized to magnesium ion, while copper (Cu) in aqueous solution is being reduced to copper solid.
- Looking at the standard reduction potentials table for each element we find the following:

Mg^{2+} (aq) + 2 e^- → Mg (s) E^0 = -2.37 volts

Cu^{2+} (aq) + 2 e^- → Cu (s) E^0 = +0.34 volts

- Now, you need to change the SRP to match the actual reaction. In the above reaction, magnesium is oxidized and copper is reduced. Therefore:

Ox ½ reaction: Mg (s) → Mg^{2+} (aq) + 2 e^- E^0 = +2.37 volts*

Red ½ reaction: Cu^{2+} (aq) + 2 e^- → Cu (s) E^0 = +0.34 volts

*notice the sign for the potential difference has changed

- Combine the potentials to get a net potential for the overall reaction.

+2.37 V + +0.34 V = +2.71 V

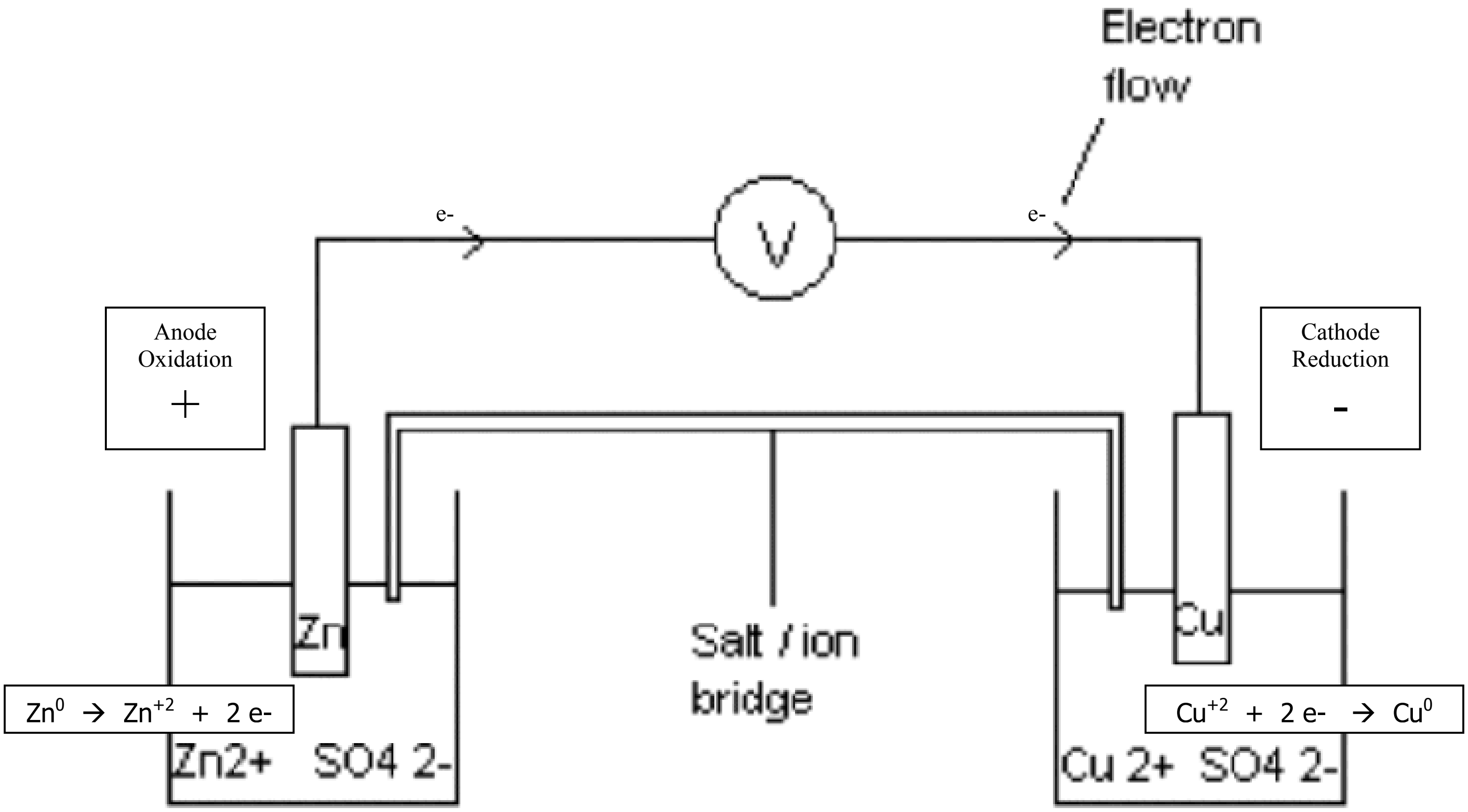
- This represents the theoretical voltage produced in this electrochemical cell.

2. The above redox reaction is very simple since we are assuming that there are one mole of each reactant and product. Plus, it just so happens that there are 2 moles of electrons produced in each half reaction.

3. It is wise to provide students with reinforcement exercises on determining half reactions, balancing half reactions, calculating net potential in a redox reaction.

Electrochemical Cells:
A characteristic of electrochemical cells is that the redox reaction may occur spontaneously (voltaic cell), or non-spontaneous reactions can be forced to occur (electrolytic cell & electroplating).

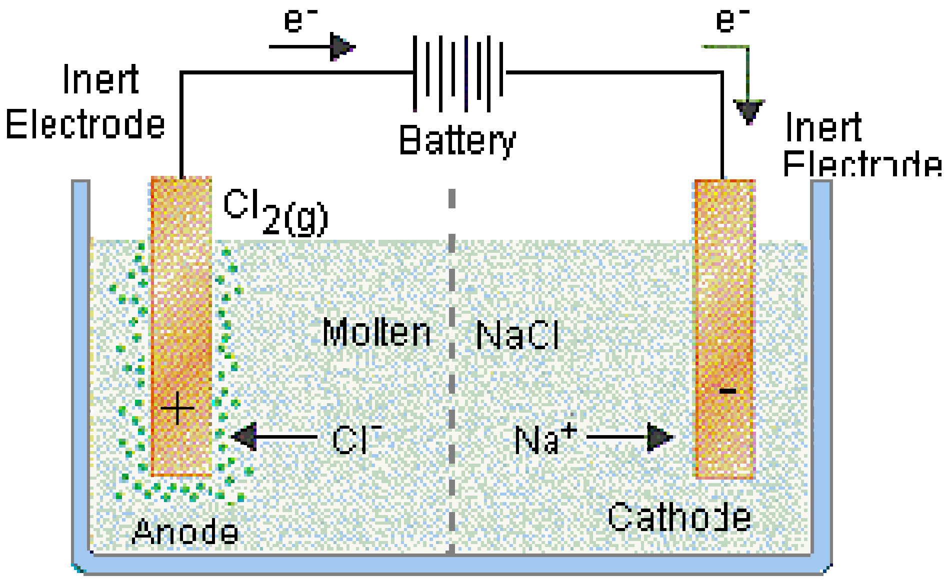
1. **Voltaic cells** are typically used to produce electrical energy. In fact, batteries are a voltaic cell.
2. **Electrolytic cells** use electricity to bring about a redox reaction that would normally be non-spontaneous. In other words, low energy reactants become high energy products.
3. **Electroplating** is a process that coats a material with a layer of metal based on metallic activity.
4. **Rechargeable batteries** - Spontaneous redox reactions eventually deplete the electrons available at the anode (e.g. causing a battery to become "dead"). The redox reactions can be "reversed" by a non-spontaneous reaction.
   a. Voltaic cells are "batteries" that supply potential (voltage or electrical energy) to run motors, light bulbs, etc.
   b. Eventually, the potential in the voltaic cells will deplete because the electrons from the anode (oxidation) are depleted and built up at the cathode (reduction).
   c. A batteries or outside voltage supply can be connected to the voltaic cell so that the anode is connected to the negative terminal (supplying electrons that were lost) and the cathode is connected to the positive terminal (removing the electrons that were gained).



Voltaic cells

a. In a voltaic cell, electrons are spontaneously emerging at the cathode so that reduction can occur.
   i. The electrons are "pumped" from the anode towards the cathode, supplying the resistance (bulb, motor, cell, etc.).
   ii. Since the electrons leave the anode, it is designated with a positive charge. The electrons are built up at the cathode; therefore, the cathode is often denoted with a positive sign.
   iii. Strictly speaking, the term "cathode" was derived because "cations" (positively charged ions/particles) were attracted to it. The "anode" attracts the "anions" (negatively charged ions/particles).
   iv. The salt bridge completes the circuit so that current (electrons) can flow. If the circuit is "broken", the half reactions will not occur.

b. The electrons are the source of potential in the voltaic cell.
   i. As the electrons leave the anode (Zn), the Zn metal disintegrates.
   ii. As the electrons reduce at the cathode (Cu), the Cu metal increases in mass at the cathode.



Electrolytic cells

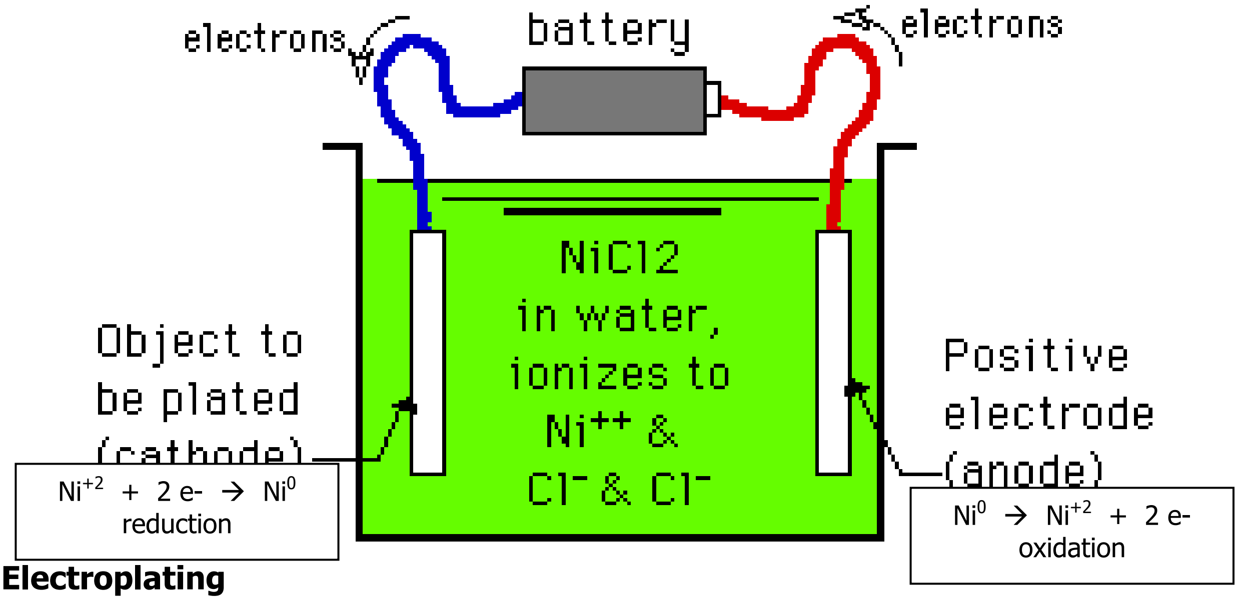
a. The example above shows why the process is called electrolysis. The suffix -*lysis* comes from the Greek stem meaning to loosen or split up. Electrolysis literally uses an electric current to split a compound into its elements, including reactions such as the electrolysis of brine (NaCl into sodium and chorine gas), electrolysis of aluminum oxide to aluminum and oxygen gas, and the electrolysis of water into hydrogen and oxygen gases.

b. Non-spontaneous: In an electrolytic cell, the electrons are being forced to move in a direction opposite of a spontaneous reaction.

c. The cathode where reduction occurs is often denoted with a negative sign. When Na^+ ions collide with the negative electrode, the battery carries a large enough potential to force these ions to pick up electrons to form sodium metal. The reduction half reaction is: $Na^+ + e^- \rightarrow Na$.

d. Since electrons are being removed from the anode after oxidation, it is given a positive sign. <u>Anions</u> (negative chlorine ions above) are attracted to the anode. Cl^- ions that collide with the positive electrode are oxidized to Cl_2 gas, which bubbles off at this electrode. $2\ Cl^- \rightarrow Cl_2 + 2\ e^-$

Using the SRP reference table, the potential required to oxidize Cl^- ions to Cl_2 is -1.36 volts and the potential needed to reduce Na^+ ions to sodium metal is -2.71 volts. The battery used to drive this reaction must therefore have a potential of at least 4.07 volts.



Electroplating

a. Non-spontaneous: Electroplating is the deposition of a metallic coating [*nickel in this case*] on an object at the cathode by exposing it to a solution containing a salt [*e.g. nickel chloride in this case*] of the metal to be deposited and then, putting a negative charge on the object, whereby positively charged metal ions [*Ni^{+2} ions in this case*] in the salt solution are reduced to the metallic element form.

b. The $NiCl_2$ salt ionizes in water into Ni^{+2} ions and two parts of Cl^- ions.

c. The object to be plated is negatively charged (*by being connected to the negative pole of the battery*). Therefore, it attracts the positively charged Ni^{+2} ions.

d. The Ni^{+2} ions reach the object to be plated, and electrons flow from the object to the Ni^{+2} ions.

e. For each ion of Ni^{+2}, 2 electrons are required to neutralize its positive charge and "reduce" it to a metallic atom of Ni^0.

Common Electroplating Experiments:

1. Zinc plating pennies

   - The cathode will be copper (pennies), the anode will be zinc, and the electrolyte (solution) will be zinc dissolved in vinegar and water. (Zinc anodes are available from boating stores)

2. Copper plating a key or a quarter

   - The cathode will be brass (key/quarter), the anode will be copper, and the electrolyte (solution) will be copper sulfate with sulfuric acid dissolved in water.

Terms and Concepts

Anode	Half-reactions	Reduction
Cathode	Oxidation	Redox
Electrochemical Cell	Oxidation State	Salt Bridge
Electrolysis	Oxidizing Agent	Spontaneous
Electroplating	Reducing Agent	SRP

Instructional Resources

Sources for background information

- Notes from Craig T. Riesen, Clarenceville High School
- http://www.wwnorton.com/college/chemistry
- www.google.com sites for electrochemical cell, electrolysis, electroplating, and Standard Reduction Potential reference table.

Print resources

- Notes from Craig T. Riesen, Clarenceville High School
- Natural Approach to Chemistry, Lab Aids Curriculum (Hsu, Chaniotakis, Damelin)

Suggested online resources for instruction and/or teacher background

- "moodle" folder from Craig T. Riesen, Clarenceville High School: www.clarenceville.k12.mi.us "moodle" "online learning" "science" "Chemistry A or B" … the enrolment key is the course title
- http://www.wwnorton.com/college/chemistry

Electrochemistry (Voltaic Cells)

Activity #1 – The Lemon Battery Part 1

Questions to be investigated

What kind of chemistry exists in a battery?

Objectives

- Students discover that a lemon can create electricity (as in a household battery) based on a chemical reaction (redox reaction)

Teacher Notes

- This activity is intended to be sequential (followed by Activities 2 - 3). It should "hook" student interest for the topic of electrochemistry. It can be used to introduce the unit or as a reflective exercise.
- Activity #2 introduces the concept of voltage being the "potential difference" produced between chemical reactions.
- Activity #2A uses the lemon battery to show the principle of rechargeable batteries.
- Activity #4 causes students to realize that one type of chemical reaction involving potential difference is the reduction / oxidation (redox) reaction.
- An enrichment exercise: try using a lime, olive, etc. rather than a lemon.

Materials

One lemon	One light emitting diode (LED)
5 cm piece of copper wire	5 cm piece of magnesium ribbon

Two alligator clips → using two different colored alligator clips is very helpful

Safety Concerns

Normal laboratory safety procedures should be used.

Real-World Connections

Everything on earth stores and uses energy to stay alive. The ultimate source of energy is the sun. Basically this means that the sun produces all the energy used in our lives on planet earth. However, we also produce and store energy. By what process does life collect and store its energy? [Give students a chance to respond.] In this experiment, students will see how the process of energy storage (and use) in living things is described by "electrochemistry." Our bodies, and the lemon used in this activity, are both sources of energy that can be stored or used to supply energy to operate something else. In other words, we and the lemon are both living batteries!

Sources

- Craig T. Riesen
- Natural Approach to Chemistry, Lab Aids Curriculum (Hsu, Chaniotakis, Damelin)

Procedure/Description of Lesson

1. Roll the lemon on the table using your palm so that the juice can flow easily.
2. Make two 1 cm incisions in the lemon about 2 cm apart from one another.
3. Insert the magnesium ribbon piece in one of the incision holes.
4. Insert the copper wire in the other incision hole.
5. Attach one of the alligator clips to the magnesium ribbon and to one side of the LED.
6. Attach another alligator clip to the bare end of the copper wire and to the other side of the LED.
    - If the LED does not light, switch the way that the LED leads are connected to the copper wire and the magnesium ribbon.

Discussion Questions

1. What happened to the LED when the lemon "battery" was connected properly?

2. What must be produced from the lemon to light up the LED?

3. Where did the electricity come from?

Electrochemistry

Activity #2A – The Lemon Battery Part 2

Questions to be investigated

What makes a lemon battery work (light up a bulb)?

In a system, what must be present for electricity to flow or for voltage to be produced?

Objectives

- Students will learn that voltage is the "potential difference" between two reactions.
- Students will understand that two chemical reactions takes place to make the lemon battery operate to light the LED or to give a voltage.

Teacher Notes

This Activity can be a direct follow up to Activity #1 or a stand alone reflective exercise. (See Teacher Notes for Activity #1)

Materials

One lemon "Chemistry Probe" or voltmeter (*measures 1-3 V*)
5 cm piece of copper wire 5 cm piece of magnesium ribbon
Two alligator clips → using two different colored alligator clips is very helpful
One light emitting diode (LED)

Safety Concerns *

Normal laboratory safety procedures should be used.

Sources

- Craig T. Riesen
- Natural Approach to Chemistry, Lab Aids Curriculum (Hsu, Chaniotakis, Damelin)

Procedure/Description of Lesson

1. Use the same set up as Activity #1 initially so that the LED lights up.
2. Have students reverse the alligator clips on the LED. Does anything happen? (**certain LED lights are "one way" and that kind would not light up, showing that the flow of charge is in a particular direction**)
3. Have students disconnect one of the alligator clips.
   a. What happens to the LED? (does not light up)
   b. Why doesn't the LED light up … in other words, what is necessary for the lemon battery to work? (**circuit … complete loop of charge**)
4. Have students disconnect the LED completely and put it aside.
5. Turn on the voltmeter and notice the "0.0 V" reading. Zero volts represents the starting point at which nothing is happening chemically. When a voltage reading changes from zero, this indicates chemical activity.
6. Connect the alligator clips from the magnesium ribbon to one of the terminals of the voltmeter, and the alligator clip from the copper wire to the opposite terminal of the voltmeter.
   a. Allow 30 seconds for the voltmeter to establish a stable reading.
   b. What is the voltage of the lemon battery? Be sure to label whether the voltmeter shows a positive or negative voltage. (ex. +1.9 V or -1.9 V)
7. Switch the way that the LED leads are connected to the copper wire and the magnesium ribbon
   a. Allow 30 seconds for the voltmeter to establish a stable reading.
   b. What is the new voltage of the lemon battery? Be sure to label whether the voltmeter shows a positive or negative voltage. (ex. +1.9 V or -1.9 V)

Analysis Questions

1. Student should answer the questions embedded in the procedures before completing this section.
2. Before "connecting" the lemon battery, the voltage should be "0.0 V." After connecting the lemon battery, a non-zero voltage is produced. How is voltage determined within the voltmeter in terms of mathematics?
3. What procedure in the activity caused the voltage to change from + to – or – to +?
4. Explain why the voltage (potential difference) changed from + to – or – to +.
5. Formative Assessment:
   a. Use the analog of a waterfall to define or describe "potential difference" or voltage
   b. What kind of energy interactions are involved in this activity?

Answers to Questions

1. Student should answer the questions embedded in the procedures before completing this section.
2. Before "connecting" the lemon battery, the voltage should be "0.0 V." After connecting the lemon battery, a non-zero voltage is produced. How is voltage determined within the voltmeter in terms of mathematics? (**Since voltage actually means "potential" in terms of electrical energy, a DIFFERENCE in POTENTIAL must exist for voltage to change**. The voltage or potential changes from zero to a new potential based on the difference in the reactions present.)
3. What procedure in the activity caused the voltage to change from + to – or – to +? (**When the electrodes are switched on the voltmeter.**)
4. Explain why the voltage (potential difference) changed from + to – or – to +. (**There must be two opposing chemical reactions taking place since a positive voltage and a negative voltage are produced when the electrodes are reversed.**)
5. Formative Assessment:

   a. Use the analog of a waterfall to define or describe "potential difference" or voltage**. (A waterfall works based on the fact that there is a difference is height (potential to fall). The greater the height the water falls, the greater the potential. This equates to voltage: the greater the potential difference between two reactions, the more voltage that is present. If water is on the level, there will be no flow (no potential difference.)**

   b. What kind of energy interactions are involved in this activity? **(potential energy → voltage; kinetic energy → flow of electrons or production of electricity to do work (light the LED).)**

Electrochemistry

Activity #2B – Rechargeable Batteries (Lemon Battery Part 3)

Questions to be investigated

- Can chemical reactions be sustained indefinitely?
- Can anything be done about a system (a battery in this case) whose charge is depleted?
- If life hands you lemons, can you make lemonade, I mean, a battery that is useful?

Objectives

- Students will understand that many chemical reactions are reversible and that this is seen in our daily lives.
- Students will learn that a spontaneous chemical reaction (lemon battery) will eventually "run out" or be depleted and this reaction can be "restored" by applying energy (in this case voltage).

Teacher Notes

Activity #3 logically follows and should follow Activities 1 & 2 after the lemon battery has "run" for a period of time.

Materials

One lemon "Chemistry Probe" or voltmeter (*measures 1-3 V*)
5 cm piece of copper wire 5 cm piece of magnesium ribbon
Two alligator clips → using two different colored alligator clips is very helpful
One light emitting diode (LED) 6 Volt Battery

Safety Concerns

Normal laboratory safety procedures should be used.

Real-World Connections

So much of our lives involve reversible chemical reactions. Even nature itself uses reversible "cycles" (water cycle, nitrogen cycle, oxygen cycle, etc.) demonstrating that substances are used and renewed based on reversible reactions (precipitation vs. evaporation). We use the principle of reversible reactions to refrigerate our food or to air condition our homes and vehicles (for every degree we cool down we must warm up a degree. NiCd batteries are used in many rechargeable devices (cordless phones, razors, cameras, and power tools).

Sources

- Craig T. Riesen
- Natural Approach to Chemistry, Lab Aids Curriculum (Hsu, Chaniotakis, Damelin)

Procedure/Description of Lesson

1. Connect the positive terminal of the 6 V battery to the former "anode" (the electrode where oxidation took place and lost electrons) using one of the alligator clips. (**the magnesium ribbon**)
2. Connect the negative terminal of the 6 V battery to the former "cathode" (the electrode where oxidation took place and lost electrons) using the other alligator clip. (**the copper wire/strip**)
3. Allow the system to sit for several minutes and make observations.

Assessment Ideas

1. Formative assessment: How can a chemical reaction be reversed?

2. List some practical, daily, or industrial applications developed from reversible reactions?

3. Explain the "recharging" process.

Answers to Questions

1. Formative assessment: How can a chemical reaction be reversed? (**a spontaneous reaction will eventually "run out." However, using an outside energy source can reverse the reaction and replenish the supply of chemicals for the spontaneous reaction to proceed**.)

2. List some practical, daily, or industrial applications developed from reversible reactions? (**rechargeable batteries, downhill skiing with the ski lift, clock reactions, body reactions like CO_2 and HCO_3^-, water cycle in nature, refrigeration, oxygenation/deoxygenation of blood, nerve impulses sent to and from the brain as in a reflex reaction, etc.**)

3. Explain the "recharging" process. (**During oxidation the electrons from magnesium are lost (used up). By connecting to the negative terminal of the 6 V battery, the electrons (from the battery) can replenish to some degree the electrons lost by the magnesium.**)

Electrochemistry

Activity #3A – Electroplating

Questions to be investigated

What is taking place in electrochemistry?

How is electroplating practical to us?

Objectives

Students will learn that two chemical reactions take place in electrochemistry.

Teacher Notes

- This Activity can be a direct follow up to Activity #2A or a stand alone reflective exercise. (See Teacher Notes for Activity #1)
- Activity #3A leads students to see electroplating (which is cool) and that two chemical reactions are occurring.
- Activity #3B leads students to recognizing the reduction and oxidation (the two chemical reactions) chemical reactions taking place during electroplating.
- You could run both Activity #3A and 3B simultaneously (split the class in half) to save time.

Materials

6 V Battery	0.5 M Sulfuric Acid (weak electrolyte)
Strip of Copper Metal	Strip of Silver or brass, a key, quarter
1 M $CuSO_4$ solution	250 ml beaker

Two alligator clips → using two different colored alligator clips is very helpful

Safety Concerns

Normal laboratory safety procedures should be used.

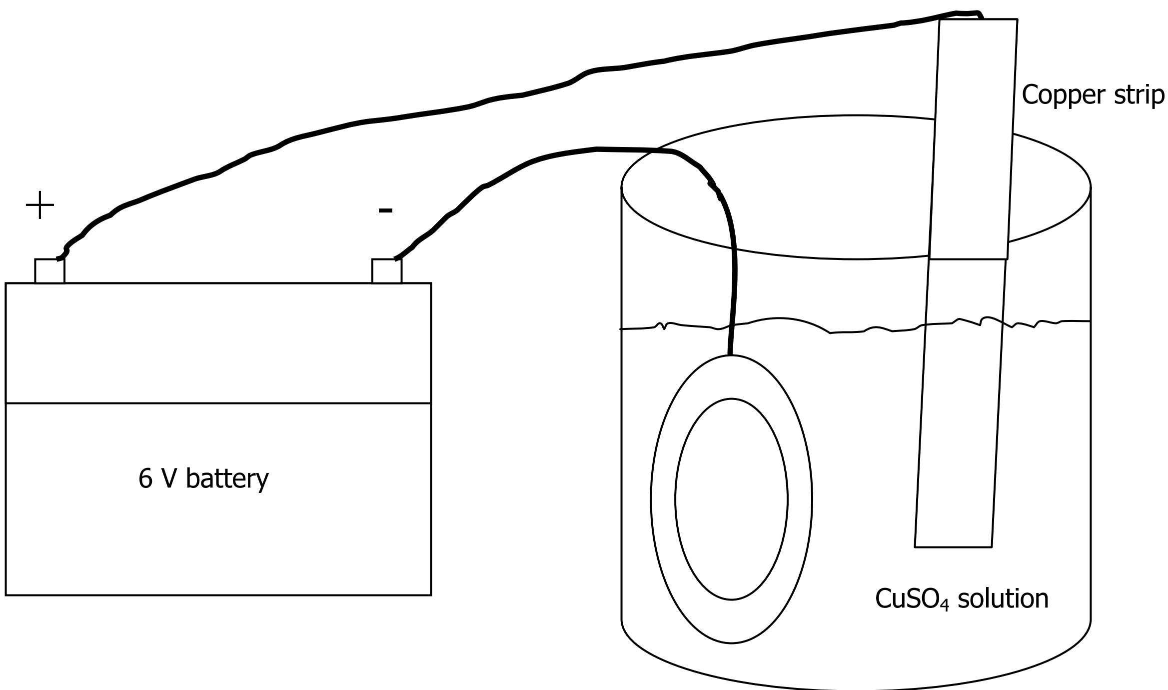
Real-World Connections

Countless items in our lives are electroplated. Rings, necklaces, bracelets, and various metal covers are often electroplated. Manufacturers take a cheap metal and plate a very thin layer of an expensive metal (gold, silver, etc.) on the outside. The appearance looks as if the entire substance is pure. Chrome plating is used to weather proof metals. Vehicles are painted today using electroplating which prevents future rust. Many metals are mined and refined using electroplating.

Sources

Craig Riesen & Traci Banjanin.

Procedure/Description of Lesson



1. The procedural set up is shown above.
2. Pour 150 ml of copper II sulfate ($CuSO_4$) solution to the 250 ml beaker.
3. Add 10 ml of the 0.5 M sulfuric acid solution to the beaker.
4. Connect the positive (+) terminal of the battery to the copper strip using an alligator clip.
5. Connect the negative (-) terminal of the battery to the item being plated (quarter, silver strip, etc.)
6. Place both the copper strip and the item being plated into the beaker solution at the same time.
7. Allow the apparatus to stand for several minutes, observing any changes or happenings (bubbles, color change, etc.). <u>Answer Assessment question 1 before going on.</u>
8. After five minutes, remove the alligator clips from the copper strip and the item being plated. Remove the metals and place them on a piece of paper towel for observation. <u>Answer assessment questions 2-4 before going on.</u>
9. Once you have answered assessment questions 1-4, you will switch the battery connections with your electrodes.
   i. Switch the alligator clip connected to the copper strip with the item being plated
   ii. Switch the alligator clip connected to the item that was plated with the copper strip.
10. Allow this system to sit for several minutes. <u>Answer assessment question 5</u>.

Assessment Ideas

1. Record your observations of the electroplating process (procedure 7) prior to observing the copper strip and item being plated.
2. Now specifically look at the item being plated. What happened to the item being plated in terms of color, material, etc.?
3. Now specifically look at the copper strip. What happened to the copper strip?
4. Based on this activity, how many chemical reactions are taking place during electroplating
5. Make observations based on procedures 9-10. What happened to the copper strip, the electroplated item, and the salt solution?
6. Formative Assessment:
    a. Have students walk around the room; go through their pockets, share ideas and to list as many items as they can find that may be electroplated.
    b. Take one item they found and explain the process of electroplating. In other words, what metal and salt solution was used to plate the item?
    c. Thirty years ago, cars in Michigan would rust within two years. Within five years "body work" would need to be done. Today, vehicles can last years before a small amount of rust forms. What makes the difference
    d. Look up various metals and research how they are made available to us (e.g. copper, zinc)

Answers to Questions

1. Record your observations of the electroplating process (procedure 7) prior to observing the copper strip and item being plated.
2. Now specifically look at the item being plated. What happened to the item being plated in terms of color, material, etc.? (**There should be a greenish copper color on the item**)
3. Now specifically look at the copper strip. What happened to the copper strip? (**If enough time passed, the copper strip could be slightly mottled or pitted, indicating that some of the copper was removed in the chemical reaction**.)
4. Based on this activity, how many chemical reactions are taking place during electroplating? (**TWO: one at the copper strip (depleting it) and the other at the item being plated because it is turning a copper color**)

5. Make observations based on procedures 9-10. What happened to the copper strip, the electroplated item, and the salt solution?

6. Formative Assessment:

    a. Have students walk around the room; go through their pockets, share ideas and to list as many items as they can find that may be electroplated.

    b. Take one item they found and explain the process of electroplating. In other words, what metal and salt solution was used to plate the item?

    c. Thirty years ago, cars in Michigan would rust within two years. Within five years "body work" would need to be done. Today, vehicles can last years before a small amount of rust forms. What makes the difference? (**coat hangers**, **chrome plating, Car paint primer uses a similar process to electroplating to deposit a polymer onto a metal and prevents rust from forming)**

    d. Look up various metals and research how they are made available to us (e.g. copper, zinc) (**copper and some other metal refining involves electroplating**)

Electrochemistry

Activity #3B – Electroplating with Silver Nitrate

Questions to be investigated

Can you recognize and define two chemical reactions taking place in electrochemistry?

Objectives

Students will discover and understand reduction and oxidation half reactions that take place in electrochemistry.

Teacher Notes

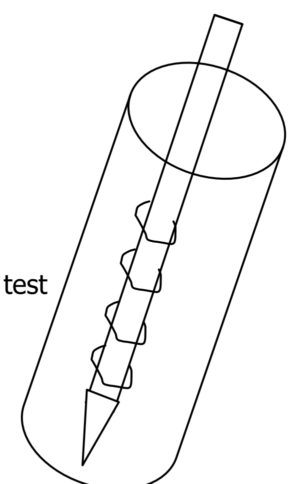
- See Part 3A notes.
- Prepare 100 - 200 ml of the 1 M $AgNO_{3\ (aq)}$ solution and have groups use only 10 - 20 ml each. (It is expensive!).
- As an addendum, once students have completed Activity 3B, have them look at the pencil lead. It should also get silver plated. You can then reflect and review the same process, leading them to discover the reduction and oxidation reactions.

Materials

Bare Copper Wire	1 M $AgNO_{3\ (aq)}$	Test Tube Rack
Large Test Tube	Pencil with sharpened lead	

Procedure/Description of Lesson

1. The sketch at the right will be used.
2. Sharpen a pencil so that the lead is around 1 cm long.
3. Wrap 10-15 cm of bare copper wire around the lead end of the pencil.
4. Pour $AgNO_{3\ (aq)}$ solution into the test tube until it is about 2/3 full.



5. Allow the pencil with the copper wiring to sit for a minute or so and then <u>record observations</u> regarding the $AgNO_{3\ (aq)}$ solution and the bare copper wiring as assessment 1.

Assessment Ideas

1. Record your observations of the electroplating process (procedures 1-5) prior to observing the copper strip and item being plated.

2.

   a. In your lab group, determine why the copper metal turned silvery (actually, silver is deposited on the copper wire).

   b. Write a simple chemical equation showing what happened. Write out the elements with their oxidation state, showing what happened:

   c. Write a specific chemical equation, showing what happened.

   d. What happened to the oxidation state of the copper and why?

   e. What is the name of the chemical reaction that took silver ions out of solution and deposited on the copper wire as the pure element, silver?

3.

   a. In your lab group, determine why the $AgNO_{3\ (aq)}$ solution turned blue.

   b. Write a simple chemical equation showing what happened. Write out the elements with their oxidation state, showing what happened:

   c. Write a specific chemical equation, showing what happened.

   d. What happened to the oxidation state of the copper and why?

   e. What is the name of the chemical reaction that dissolved the pure copper element causing ions to form in solution?

4. State the two chemical reactions taking place in electrochemistry. Define each in terms of electrons gained or lost.

Answers to Questions

1. Record your observations of the electroplating process (procedures 1-5) prior to observing the copper strip and item being plated. (**Students should notice that the $AgNO_{3\,(aq)}$ solution is turning bluish (copper) and the copper wiring is turning silverfish**.)

2. In your lab group, determine why the copper metal turned silvery (actually, silver is deposited on the copper wire).

   a. (**Students need to recognize that Ag^{+2} ions from the solution deposited onto the copper wire yielding pure silver (Ag^0)**.)

   b. Write a simple chemical equation showing what happened. Write out the elements with their oxidation state, showing what happened:

   $$Ag^{+2} \rightarrow Ag^0$$

   c. Write a specific chemical equation, showing what happened.

   $$Ag^{+2} + 2\,e- \rightarrow Ag^0$$

   d. What happened to the oxidation state of the copper and why? (**it decreased, was "reduced" because electrons were gained**)

   e. What is the name of the chemical reaction that took silver ions out of solution and deposited on the copper wire as the pure element, silver? (**REDUCTION**)

3. In your lab group, determine why the $AgNO_{3\,(aq)}$ solution turned blue.

   a. (**Students need to recognize that Cu^{+2} ions went into solution from the copper wire which is pure copper (Cu^0)**.)

   b. Write a simple chemical equation showing what happened. Write out the elements with their oxidation state, showing what happened:

   $$Cu^0 \rightarrow Cu^{+2}$$

   c. Write a specific chemical equation, showing what happened.

   $$Cu^0 \rightarrow Cu^{+2} + 2\,e-$$

   d. What happened to the oxidation state of the copper and why? (**it increased because electrons were lost**)

   e. What is the name of the chemical reaction that dissolved the pure copper element causing ions to form in solution? (**OXIDATION**)

4. State the two chemical reactions taking place in electrochemistry. Define each in terms of electrons gained or lost. (**reduction** gains electrons and **oxidation** loses electrons)

Electrochemistry

Activity #4 – Squeezing out the Last Drop of the Lemon Battery

Questions to be investigated

- What are the Standard Reduction Potentials for the reduction and oxidation half reactions taking place in the lemon battery?
- Are all redox reactions created equal? In other words, what happens to the battery voltage (potential difference) if a difference fruit is used in place of the lemon?

Objectives

- Students will write reduction and oxidation half reactions.
- Students will calculate the voltage of each half reactions as well as the total redox voltage using the SRP table.
- Students will compare a different battery using an olive rather than a lemon (still using the same electrodes: magnesium and copper).

Teacher Notes

- You do not even have to set up the Activity again, but can simple have students complete the activity mathematically based on their previous results.
- If students are struggling with the concepts, use the exact same set up in Activity #2A and determine the voltage of the reaction.
- It is important to note whether the voltage is + or – because that shows which reaction (magnesium or copper) is being reduced and oxidized.

Materials

One lemon	"Chemistry Probe" or voltmeter (*measures 1-3 V*)
5 cm piece of copper wire	5 cm piece of magnesium ribbon
Two alligator clips	One light emitting diode (LED)
Olives and/or Limes	SRP table

Sources

- Craig T. Riesen
- Natural Approach to Chemistry, Lab Aids Curriculum (Hsu, Chaniotakis, Damelin)

Safety Concerns

Normal laboratory safety procedures should be used.

Procedure/Description of Lesson

1. Refer to the Lemon Battery Activity #2A,
   a. Write the reduction and oxidation half reactions. (based on magnesium and copper reacting in the lemon electrochemical cell)
   b. Write the theoretical half reaction potentials (based on the electrodes) and determine the theoretical voltage (potential difference) of the electrochemical cell (use the SRP table).
   c. Calculate the percent error for your experiment results (Activity 2A voltage) with the theoretical voltage you found in 1a.
2. Use an olive or a lime and repeat Activity 2A to determine the voltage (potential difference) in the "electrochemical cell."
   a. Use your theoretical calculations from procedure 1b to determine the percent error for your experimental results.
   b. How does the olive battery compare with the lemon battery in terms of potential?

Assessment Ideas

1. If you are given the following materials: magnesium ribbon, copper wire, a lemon, alligator clips, voltmeter; how would you know which metal is reduced and which is oxidized BEFORE completing the activity? In other words, predict the element that will be reduced and which will be oxidized.

   (Use the SRP table to determine which metal is most active … that will be the reducing agent (oxidized). In the lemon battery, if the voltage is negative, then magnesium is oxidized since it is the more active metal. If the voltage is positive, think the opposite, copper is oxidized)

2. Using the law of conservation of mass and charge in order to make a statement about the relationship between reduction and oxidation half reactions.

3. Which element in this activity is the oxidizing agent and which is the reducing agent? Use specific data and measurements with your answer because the potential difference, the sign (+ or -) are factors to incorporate.

4. The following equation reaction represents a chemical cell at 298 K. When the electrochemical cell is allowed to spontaneously operate, what would the maximum potential (E^0) for the cell be?

   $$2\ Al^0\ (s) + 3\ Cu^{2+}\ (aq) \rightarrow 2\ Al^{+3}\ (aq) + 3\ Cu^0\ (s)$$

   *Notice that Al is oxidized so the SRP sign must be opposite.

Electrochemistry

Activity #5 – Electrochemical Cells: The Battery

Questions to be investigated

- How do batteries store energy?
- Why is the energy of a battery used only when something is connected across its terminals?

Objectives

- Students will recognize reduction and oxidation using an electrochemical cell.
- This activity can be used to do basic mathematics of half reactions or used simply to build a simple electrochemical cell with a spontaneous reaction.

Teacher Notes

- It is good to remind students over and over that electrochemical cells come in two brands: spontaneous (batteries) and non-spontaneous (needing an energy supply as in electrolysis and electroplating).
- The activity uses magnesium and copper as in activity 4. Therefore, the mathematics is the same. However, in activity 4, students were not asked to come up with the overall chemical equation represented by the electrochemical cell.

Materials

1 M $MgSO_{4\,(aq)}$	LED	5 cm piece of bare copper wire
1 M $CuSO_{4\,(aq)}$	2 Alligator clips	5 cm piece of bare magnesium ribbon
2 50 ml beakers or 2 small vials		Chemistry probe system/voltmeter
3 cm copper wire with insulation (but ends bare)		

Safety Concerns

Normal laboratory safety procedures should be used.

Real-World Connections

Batteries store chemical energy and convert it to electrical energy by connecting the positive and the negative terminals. When the terminals are connected, electricity flows. The flow of electricity is a result of a chemical reaction that takes place inside the battery. The chemical reaction depends on the materials that the battery is made of.

Sources

- Craig T. Riesen
- Natural Approach to Chemistry, Lab Aids Curriculum (Hsu, Chaniotakis, Damelin)

Procedure/Description of Lesson

1. Add enough $MgSO_4$ and $CuSO_4$ solutions to the 50 ml beakers or vials to fill them 2/3 full.
2. Place the magnesium ribbon in the $MgSO_4$ solution and the bare copper wire in the $CuSO_4$ solution.
3. Connect the red alligator clip of the voltage probe to the copper wire and the black alligator clip of the voltage probe to the magnesium ribbon.
4. Turn on the chemistry probe system and measure the voltage. Answer assessment question 1 before going on.
5. Now take the 3 cm piece of insulated copper (with bare ends) and bend it into a "U" shape. Place this as a "bridge" connecting the two beakers or vials.
6. Turn on the chemistry probe system and measure the voltage. Answer assessment question 2 before going on.
7. Remove the copper wire "bridge" and notice the voltage measurement on the probe. It should be zero again, proving that voltage in an electrochemical cell needs a completed circuit.
8. Replace the copper bridge and allow the apparatus to sit for 3-5 minutes. Then, remove the two electrodes and look at the ends that were in the solutions. Answer assessment question 3-5.

Assessment Ideas

1. What voltage did you measure in procedures 1-4? Explain why this value makes sense.

2. What voltage did you measure in procedures 5-6? What is necessary for a non-zero voltage reading (what does the "bridge" produce that was not previously present)?

3. Record your observations of the magnesium ribbon electrode, the solutions, and the copper wire electrode. Notice the color, whether a deposit forms, and whether a depletion of one of the metals can be observed.

4. Write the half reactions of the electrochemical cell in this activity, showing their E^0 values. Then, write the over equation represented by the electrochemical cell and calculate the overall E^0 (use the SRP table).

   *Notice that Mg is oxidized so the SRP sign must be opposite.

5. Without experimenting (because we will do that in activity 6), assume that one uses zinc and copper as the electrodes in an electrochemical cell. Write the expected half reactions, showing their E^0 values. Then, write the overall chemical equation represented by the electrochemical cell and calculate the overall E^0 (use the SRP table).

   *Notice that Mg is oxidized so the SRP sign must be opposite.

Assessment Ideas

1. What voltage did you measure in procedures 1-4? Explain why this value makes sense.
   (**The circuit is not complete because no "bridge" has been made yet, therefore, the voltage is zero**.)
2. What voltage did you measure in procedures 5-6? What is necessary for a non-zero voltage reading (what does the "bridge" produce that was not previously present)?
   (**A potential difference must exist for voltage to be produced. When the "bridge" is absent, there is no potential difference as the voltage on both sides is zero. However, when the "bridge" is present, the two sides possess different voltages (potentials) and therefore, yield a potential difference**.)
3. Record your observations of the magnesium ribbon electrode, the solutions, and the copper wire electrode. Notice the color, whether a deposit forms, and whether a depletion of one of the metals can be observed.
   (**The magnesium ribbon begins to be depleted (mottled, pitted, etc.) while the copper strip should have deposits at it base.**
4. Write the half reactions of the electrochemical cell in this activity, showing their E^0 values. Then, write the over equation represented by the electrochemical cell and calculate the overall E^0 (use the SRP table).

Ox ½ reaction: $Mg\ (s) \rightarrow Mg^{2+}\ (aq) + 2\ e^-$ $E^0 = +2.37$ **volts***

Red ½ reaction: $Cu^{2+}\ (aq) + 2\ e^- \rightarrow Cu\ (s)$ $E^0 = +0.34$ **volts**

$$Mg\ (s) + Cu^{2+}\ (aq) \rightarrow Mg^{2+}\ (aq) + Cu\ (s)$$

$$+2.37\ V + +0.34\ V = \underline{+2.71\ V}$$

*Notice that Mg is oxidized so the SRP sign must be opposite.

5. Without experimenting (because we will do that in activity 6), assume that one uses zinc and copper as the electrodes in an electrochemical cell. Write the expected half reactions, showing their E^0 values. Then, write the overall chemical equation represented by the electrochemical cell and calculate the overall E^0 (use the SRP table).

Ox ½ reaction: $Zn\ (s) \rightarrow Zn^{2+}\ (aq) + 2\ e^-$ $E^0 = +0.76$ **volts***

Red ½ reaction: $Cu^{2+}\ (aq) + 2\ e^- \rightarrow Cu\ (s)$ $E^0 = +0.34$ **volts**

$$Zn\ (s) + Cu^{2+}\ (aq) \rightarrow Zn^{2+}\ (aq) + Cu\ (s)$$

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

*Notice that Mg is oxidized so the SRP sign must be opposite.

Electrochemistry

Activity #6 – Electrochemical Cells: The Battery

Questions to be investigated

What metal electrodes would make the strongest and most feasible battery in the classroom setting?

Objectives

Students will have a complete inquiry based experience based on electrochemistry and their previous acquired knowledge.

Teacher Notes

- This lab is completed student based and the teacher is the facilitator.
- Students will probably quickly find that lithium is the best reducing agent (lowest on the SRP table). However, you may need to have them research lithium since is extremely volatile and therefore, not practical to use in their experiment.
- Students can be directed to research the practicality of using metals for electrodes (since lithium is too volatile for the classroom). Questions should be directed to research (e.g. why are lithium batteries used throughout industry in our cell phones, portable devices, etc.?).
- It is likely that you may not have the necessary metals for the students to carry out their experimental electrochemical cell. If this is the case, use the SRP table and choose metals near the ones they chose and allow them to use those.

Materials

Students decide.

Safety Concerns

Certain metals are explosive and very volatile. Teachers must use caution as students build their battery.

Real-World Connections

Students will include this section in their experiment report.

Sources

Have students record all sources they research and use to produce their experiment.

Procedure/Description of Lesson

1. Research various metals as electrodes, checking with the teacher for viability.

2. Once the metals for the electrodes have been finalized:

   a. Write half reactions with their potentials (E^0).

   b. Write the overall electrochemical equation and determine the theoretical potential (E^0).

3. Build the actual electrochemical cell with the metal electrodes and obtain experiment results. (Possibly, the metals will need to be adjusted based on availability.)

   a. What is the actual voltage in the experiment?

   b. Calculate percent error for your experiment.

4. Write up the experiment in the formal lab report format used in class throughout the semester. (Purpose, Discussion, Materials, Procedures, Calculations & Data, Conclusions & Questions, Final Summary)

Assessment Ideas

1. Be sure to research "real life applications" and include it in your lab format as a "discussion" section (after the "Purpose" and before the "Materials" sections).

2. Obtain results from at least one other group and <u>evaluate</u> whose battery was better. Some suggestions for evaluation are below:

   a. Whose voltage was higher?

   b. Whose percent error was lower?

   c. Are the metals chosen practical to use?

   d. What are sources of error(s) in the experiment (explain percent error)?