

## UNIT II

# The Metallic World

### Unit Overview

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This unit provides an overview of both electrochemistry and basic transition metal chemistry. Oxidation-reduction reactions, also known as redox reactions, drive electrochemistry. In redox reactions, one compound gains electrons (reduction) while another one loses electrons (oxidation). The spontaneous directions of redox reactions can generate electrical current. The relative reactivities of substances toward oxidation or reduction can promote or prevent processes from happening. In addition, redox reactions can be forced to run in their non-spontaneous direction in order to purify a sample, re-set a system, such as in a rechargeable battery, or deposit a coating of another substance on a surface.

The unit also explores transition metal chemistry, both in comparison to principles of main-group chemistry (e.g., the octet rule) and through various examples of inorganic and bioinorganic compounds.

### Learning Objectives and Applicable Standards

Participants will be able to:

1. Describe the difference between spontaneous and nonspontaneous redox processes in terms of both cell EMF (electromotive force) and DG.
2. Recognize everyday applications of spontaneous redox processes as well as applications that depend on the forcing of a process to run in the non-spontaneous direction.
3. Understand the basics of physiological redox processes, and recognize some of the enzymes that facilitate electron transfer.
4. Compare basic transition-metal chemistry to main-group chemistry in terms of how ions are formed and the different types of bonding in metal complexes.

### Key Concepts and People

1. **Redox Reactions:** Redox reactions can be analyzed systematically for how many electrons are transferred, whether or not the reaction happens spontaneously, and how much energy can be transferred or is required in the process.

- The Activity Series:** The activity series is a list of metals ranked by how easily a metal can be oxidized (lose electrons); the higher the metal is on the list, the more easily it is oxidized. This can be used to predict if a simple redox reaction will happen spontaneously.
- Half-reactions:** Half-reactions can be written to illustrate more clearly the reduction and oxidation processes in a redox reaction. In a half-reaction, either the oxidation or the reduction reaction is described and the electrons are explicitly included.
- Standard Reduction Potentials:** Standard reduction potentials are used to determine the potential difference between two half-reactions. This is the basis not only for determining whether the reaction will run in the first place but also for determining the work that can be done by the reaction.
- Batteries and Galvanic Cells:** Batteries are arguably the most recognizable applications of redox reactions. Non-standard conditions (such as concentrations different from 1.0 M for solution-phase species) influence the observed potential difference between half-cells in a galvanic cell. Since most applications of redox chemistry actually take place under non-standard conditions, the influence of the conditions is an important factor in designing applications that use redox reactions.
- Electrolytic Reactions:** Processes that represent the nonspontaneous direction of a redox reaction are referred to as “electrolytic.” Electrolytic processes will not run without the application of a voltage larger than that predicted for the reaction, but as long as that condition is met, the process is governed by stoichiometry.
- Transition Metals:** Transition metals share many patterns of reactivity with main-group elements, but also can react in unique ways due to the *d* electrons in their valence shell.
- Colorful Chemistry:** Many transition metals are intensely colored. When ligands bond to transition metals, the metals’ *d* orbitals split into at least two different energy levels. What makes transition metal complexes colored is that these gaps between the sets of *d* orbitals have energies that fall in the visible region of light.
- Metal Biochemistry:** Although metals make up a very small percentage of elements in physiological systems, they play some very specific and critical roles in a variety of processes. In addition to hemoglobin, a number of enzymes contain various metals; many of these enzymes facilitate steps in redox processes.

## Video

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Metals allow the transfer of electrons through a process called oxidation-reduction, or “redox,” when one species gains electrons while another loses them. Chemists take advantage of this

process by using electron transfers to power the batteries in our flashlights, phones, or cars. In biochemistry, trace metals, such as cobalt in Vitamin B12, often drive chemical reactions that are essential for human health. Redox reactions also occur without metals, as is the case when lightening hair color.

## VIDEO SEGMENT DESCRIPTIONS

### Host Science Explanation

#### “The Metallic World and Electrochemistry”

Dr. Catherine Drennan of MIT introduces metals and the various things they can do, and defines redox.

### Real World Application

#### “Powered by Redox”

A research group at Worcester Polytechnic Institute is programming high-functioning robots to assist people in basic activities. Because these robots need to be very mobile, they are powered by batteries—and the performance of these batteries is based on oxidation-reduction chemistry.

### Laboratory Demonstration

#### “Oxidation-Reduction Reaction”

The reaction of copper ions with zinc metal is a single-displacement reaction, one of the simplest forms of a redox reaction. It can be performed in a single beaker, which produces a copper film on the surface of the zinc metal.

### Laboratory Demonstration

#### “Redox: Creating an Electric Current”

Redox reactions can be set up into two “half-reactions” in separate beakers. Connection of the beakers by a salt bridge and a wire generates current through the wire. This is the basic setup of the Daniell cell, the classic spontaneous redox reaction and the basis of an early battery.

### Current Chemistry Research

#### “Vitamin B12”

Metals are found in the body only in trace amounts, but these trace amounts play crucial physiological roles. Dr. Catherine Drennan’s research group studies the chemistry of vitamin B12, specifically the reactivity of the cobalt ion in the center of the protein. The determination of which oxidation state,  $\text{Co}^{+1}$ ,  $\text{Co}^{+2}$ , or  $\text{Co}^{+3}$ , is present can be as simple as a test that takes advantage of a common transition-metal phenomenon: the different oxidation states have different colors.

### Real World Application

#### “Folic Acid in Fetal Development”

The chemistry of vitamin B12, as studied by the Drennan group, involves a reaction with vitamin B9, or folic acid. Folic acid plays a critical role in the prevention of neural tube defects in

developing fetuses. Because the development of the neural tube occurs so early in pregnancy and the folic acid must already be in the mother's system by the time she is pregnant, many commercially available foods are fortified with folic acid. This has helped to dramatically decrease the percentage of babies born in the US with neural tube defects.

### Laboratory Demonstration

#### “Non-spontaneous Redox Reaction”

Non-spontaneous redox reactions can be forced to occur by the application of an electric current. In this demonstration, an iron screw is plated with zinc metal by using the screw and a piece of zinc as electrodes in a solution containing zinc ions. This process, called “galvanizing,” is often used to make metals (especially iron) more resistant to corrosion.

### Real World Application

#### “Hair, Highlights, and Redox”

Many redox reactions involve metals, but some do not. One common cosmetic redox process is the bleaching of hair using hydrogen peroxide. The hydrogen peroxide oxidizes melanin molecules in hair, which then absorbs higher-energy wavelengths of light, causing the molecules (and the hair) to appear lighter.

## Unit Text

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### Content Overview

Redox reactions are among the simplest cases in which students can explain observed reactivity and also predict reactivity based on established patterns. Redox reactions are also the basis for a wide array of recognizable applications, and the convenient relationships between electron transfer and energy transfer allow for straightforward calculations to support how the various applications work. The text presents both the activity series for metals and the tables of standard reduction potentials in the context of predictions and quantification of reactivity, with several examples of applied redox processes.

Although the chemistry of transition metals is sufficiently rich and unique to warrant a course of its own, the behavior of transition metals is analogous in some ways to that of main group metals. Some of the classic characteristics of transition metals (their multiple stable oxidation states, their ability to form brightly-colored compounds, their ability to participate in multiple types of bonding) are introduced in comparison and contrast to the analogous main-group behaviors.

### Sidebar Content

1. **The Permanganate Ion: An Excellent Oxidizing Agent:** The permanganate ion is an example of a colorful oxidizing agent. This sidebar uses the permanganate ion to illustrate how to balance redox reactions using half reactions.
2. **pH and Redox Reactions:** This sidebar describes how pH meters use redox reactions to measure the concentration of  $H^+$  in a given solution.

3. **Galvanization:** Zinc can be electroplated onto the surface of iron, which will help prevent the iron from oxidizing (rusting). This is common practice to prevent nails and screws from rusting.
4. **Nickel Hydrazine Perchlorate:** The explosive compound nickel hydrazine perchlorate contains both an oxidizing agent and a reducing agent. If one were to simply strike it with a mortar, a violent explosion would ensue, which is exactly what happened when one graduate student was working with it. Luckily the student survived, but he ended up blowing up a corner of the lab and he lost three fingers and suffered eye damage in the process.
5. **Titanium White and Zinc Oxide:** This side bar explains why titanium white and zinc oxide appear white.
6. **Polydentate Ligands:** When one molecule donates more than one lone pair of electrons to a metal, the molecule is referred to as a “polydentate ligand,” literally meaning “having many teeth” in Latin. EDTA is a common polydentate ligand and is often used to render metals unreactive, such as when someone is suffering from heavy metal poisoning.

## Interactives

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### Historical Timeline of Chemistry

This interactive illustrates how different discoveries build upon, disprove, or reinforce previous theories. This not only reinforces basic chemistry concepts, but also emphasizes the nature of science. Scientists mentioned in the text are represented on the timeline.

### Control a Haber-Bosch Ammonia Plant Interactive

The conversion of hydrogen and nitrogen into ammonia is a nonmetallic redox reaction. While this reaction itself does not involve a metal, an iron-based catalyst helps speed up the reaction. Please note that a lesson plan and student worksheet are available online with this interactive.

## During the Session

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### Before Facilitating this Unit

Generally, we recommend watching the videos first before reading the written units. Since there are two main components for this unit—redox reactions and transition metals—it might be helpful to teach the unit in segments. The video focuses heavily on redox processes, but includes a bit of transition metal chemistry in the segment on physiological redox. The demonstrations and activities presented in this guide are also primarily redox-focused (including an example of a redox reaction that does not require a metal). The course guide includes one demonstration of the effect of ligands on color in transition metal complexes.

### Tips and Suggestions

1. **Redox processes require both oxidation and reduction.** Remind students that the half-reaction method used to balance redox equations is bookkeeping only, and that the half-reactions happen in tandem, even if the reagents are separated into half-cells as for a galvanic cell.
2. **Any reaction can be written; not all will proceed as written.** The activity series and the series of standard reduction potentials both allow for predictions of which direction a reaction will proceed spontaneously. Many students struggle with predictions of reactivity for anything more complicated than a single displacement; encourage them to write the half-reactions, find them on the table, and identify whether the reduction half-reaction as written in the equation also has the larger of the two reduction potentials on the table.
3. **d-electrons make transition metals unique.** The *d*-electrons in the valence shell of transition metals contribute heavily to the metals' patterns of reactivity, as well as to their tendency to form highly colored compounds. Emphasize that transition metals follow patterns of reactivity analogous to patterns the students have already learned, even though the specifics of transition metal reactivity are beyond the scope of this course.

### Starting the Session: Checking Prior Thinking

You might assign students a short writing assignment based on the following questions, and then spend some time discussing prior thinking. This will help elicit prior thinking and misconceptions.

1. Some metals can dissolve in acid. What do you think happens to the metal during this process?
2. Hydrogen burns readily in air. The primary product of this reaction is water, even though air contains more nitrogen than oxygen. Why do you think this happens?
3. Sodium metal reacts with water to form sodium hydroxide and hydrogen gas. How many protons and electrons does sodium have before the reaction? What about hydrogen? How many protons and electrons do the sodium and hydrogen have after the reaction?
4. Suppose you attempted to react gold, rather than sodium, with water. Would this reaction occur as vigorously?

### Before Watching the Video

Students should be given the following questions to consider while watching:

1. What makes a good battery?

2. How do biological redox reactions differ from inorganic redox reactions?
3. Can reactions be forced to run backwards?
4. Do redox reactions always have to involve metals? Why or why not?

### **Watch the Video**

#### **After Watching the Video**

The redox reactions discussed in the video include very simple (zinc with copper ions) and extremely complex redox processes (the cycle between Co(I) and Co(III) in reactions with folic acid and homocysteine). Encourage students, especially in more complex systems, to focus on the movement of electrons more than on the exact pattern of bond breakage and formation. Similarly, encourage students to view systems such as vitamin B12 in the context of the atoms immediately surrounding the metal center, rather than trying to envision the entire molecule. Use these additional questions as follow-up, either as a group discussion or as short writing assignments.

1. The Daniell cell (the copper/zinc reaction in the demonstrations) was an early battery system. What are some differences between the Daniell cell and more modern batteries, like the ones used to power the robots in the video?
2. Transition metals appear colored because of the number of *d*-electrons in the metal center and because of the ions or molecules immediately surrounding the metal. Which of these two factors is responsible for the different colors used to analyze samples in the Drennan lab? Explain why this is a useful characterization test.
3. In a beaker, a redox reaction can be forced to run in the non-spontaneous direction by applying an electric current to the components of the reaction. This is, of course, not possible in the human body. What has to happen in the body in order to force a reaction to go in the non-spontaneous direction?

## **Group Learning Activities**

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### **Dissolving a Penny from the Inside Out**

#### **Objective**

This activity will illustrate the relative redox activity of the metals used in the Daniell Cell, which is highlighted in section 5 of the Unit 11 text.

Pennies minted after 1982 are not pure copper, but rather a zinc core with a thin copper coating. Because zinc is much more reactive than copper, if a penny is scored along the edges to reveal some of the zinc, the zinc can be dissolved in acid to leave the copper behind. The reactions of the zinc core will be measurable after a few minutes, but can also be left out long enough to consume the entire core and leave the copper shell behind. Although this experiment

technically defaces US currency, it is not done for the purpose of defrauding anyone, and therefore is a mainstay demonstration and laboratory experience in general chemistry courses.

### List of Materials

- Pennies, some scored on the edge to reveal zinc core, one scored and one unscored per group
- Copper pieces
- Zinc pieces
- 6 M HCl
- Water
- Beakers, 3 per group
- Balance
- Tweezers
- Watch glasses, 3 per group

### Procedure

1. Measure and record the masses of pennies, both scored and non-scored.
2. Place both a scored and a non-scored penny into a beaker of acid. Observe and record any changes that occur.
3. While the pennies react, measure the mass of a piece of copper and the mass of a piece of zinc.
4. Place the piece of copper into a beaker of acid and the piece of zinc into a separate beaker of acid. Observe and record any changes that occur.
5. Use tweezers to remove the pennies, copper piece, and zinc piece from the acid solutions. Carefully rinse the pennies and metal pieces with distilled water and dry them on paper towels.
6. Measure the masses of the pennies and metal pieces after the reactions and compare.

### Discussion

The following questions can help students reflect on this exercise. If possible, have students record their responses to the questions in a journal and spend some time discussing as a class.

1. Which pieces of metal reacted with acid? Why do you think they reacted in that way? What does this reveal about their relative reactivities? Explain the reactions that were happening.
2. Did both pennies react when placed in acid? If not, which one did? Explain how you determined your answer.
3. Copper is a relatively unreactive metal. Why is it used only for the outside shell of pennies instead of the entire penny?



**Hazards**

It is good lab practice to review a chemical's Material Safety Data Sheet (MSDS) before working with any chemical. Follow instructions on the MSDS and encourage students to review them. 6 M HCl is highly corrosive, and the reaction with zinc may be very vigorous. Wear eye protection, gloves, and lab coats while performing this experiment. If possible, while metals are dissolving, cover beakers with a watch glass in order to prevent splashes. Use tweezers to remove any pieces of metal from beakers.

**Disposal**

Check local regulations for proper disposal of HCl.

Solids may be discarded in the trash, though unreacted copper and scored pennies are reusable for multiple lab sessions.

**“Lemon Cell Battery”**

Adapted from K. R. Muske, et al. in Journal of Chemistry Education, April 2007

**Objective**

A simple galvanic cell can be assembled using strips of different metals and a strong electrolyte solution. These cells can provide power for small electrical components, and may be connected in series to provide even more power. The lemon juice/salt solution serves as both a reagent and a salt bridge.

**List of Materials**

- Lemon juice (~3 mL per cuvette)
- Measuring spoons or graduated cylinders, one per group
- Beakers, one to three per group
- Table salt
- Stirring rods, one per group
- Funnels, one per group
- Mg ribbon, ~5 cm length per cuvette
- Cu sheet and wire, ~5 cm sheet per cuvette
- UV/vis cuvettes, one per group (more if groups are connecting cells in series)
- Double-ended alligator clip wire leads, two per group
- Single alligator clips, as many as needed to connect cells if groups are making multiple cells
- Small motor, light bulb or other electrically-powered component, or voltmeter; one per group

**Procedure**

1. Measure 3 mL of lemon juice into a beaker.
2. Stir 0.5-1.5 tsps. table salt into the lemon juice and allow to dissolve. Record the

amount used to make the solution.

3. Fold a magnesium strip over the top of one side and into the cuvette, and a copper strip over the top of the opposite side into the cuvette
4. Attach one end of an alligator clip lead to each of the strips, and attach the other end of each lead to one of the connections to a motor, LED, or other electronic component. [NOTE: if connecting multiple cells, use alligator clips to connect the magnesium side of one cuvette to the copper side of the next in the series, and connect the wire leads at the ends of the entire series.]
5. Position the motor slightly below the cuvette(s), and as far away as the wires will allow in order to prevent sparks and ignition of hydrogen from the salt solution.
6. Pour the lemon juice/salt solution into the cuvette leaving about a centimeter of the top free from the solution.
7. If using a voltmeter, record the voltage for the cell, then prepare solutions of different concentrations, replace the solutions in the cuvettes, and determine the optimum salt concentration for high cell EMF.
8. If using a motor, use the motor to power a model vehicle!

### Discussion

The following questions can help students reflect on this exercise. If possible, have students record their responses to the questions in a journal and spend some time discussing as a class.

1. If hydrogen gas is generated at the copper electrode, which electrode was the anode and which was the cathode? Write the half-reactions occurring at each electrode.
2. Based on your answers to question 1, look up the standard reduction potentials for the half-reactions and calculate the standard EMF of the cell. Compare your results with the observed EMFs if you used a voltmeter.
3. Draw a diagram of this setup and explain what is happening.

### Hazards

Hydrogen gas is generated at the copper electrode. Make sure area is well ventilated or perform under a fume hood to avoid inhalation of the gas. Fill the cuvettes only when the electrical connections have been made, and keep motors and other electrical components below the cuvettes to avoid contact with the gas. Wear eye protection, gloves, and a lab coat.

### Disposal

Lemon juice/salt solutions may be disposed in the sink. Copper and unreacted magnesium may be reused.

## In-Class Chemical Demonstrations

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### Iodine-Starch-Vitamin C Demo

Adapted from “Dr. Campbell’s Favorite Demos.”

#### Objective

This is a very quick demonstration of a redox reaction that does not involve metals. Starch forms a blue complex when combined with elemental iodine. When iodine reacts with the ascorbic acid in a vitamin C tablet, the iodine is oxidized to iodide ions (and the ascorbic acid is reduced to dehydroascorbic acid). Iodide ions do not form a complex with starch, and so the color is discharged as the iodine is oxidized. The “ingredients” in this demo can be used on a variety of surfaces (including paper!) and the vitamin C tablet can be used to discharge the color either across the entire surface or selectively.

#### List of Materials

- Tincture of iodine
- Spray starch
- vitamin C tablets from drugstore

#### Procedure

1. Swab tincture of iodine solution on paper or a watch glass.
2. Before the iodine solution dries fully, spray starch over the solution to generate a dark blue color.
3. Rub a vitamin C tablet in the iodine/starch complex to discharge the color.

#### Discussion

The following questions can help guide a discussion.

1. Starch forms a dark blue complex in the presence of iodine, but is colorless in the presence of iodide. Identify the species being oxidized and reduced in this demonstration.
2. Why is vitamin C sometimes referred to as an antioxidant?

#### Hazards

None of the components in the demo are hazardous, but tincture of iodine solution can stain just about anything. The iodine/starch complex is very goopy, so if demonstrating the reaction on your hand, wash your hands immediately afterward.

#### Disposal

All the reagents in this demo may be disposed in the trash or down the drain.

## Cobalt Painting

Adapted from *Chemical Magic* by Leonard A. Ford.

### Objective

This demonstration illustrates the change in color of cobalt (II) ions in the presence of different ligands. Both the chloride and the acetate can be dissolved in water to give a pinkish solution from the complex, but will be colored differently when the water is driven off. Solutions of the two cobalt complexes are used to paint on regular paper. When the picture is warmed near a Bunsen burner flame, the dehydration of the compounds turns blue-green and violet where the original painting was pink. If the paper is subsequently sprayed with water, the compounds rehydrate and return to pink.

### List of Materials

- Aqueous  $\text{CoCl}_2$  (~0.5 M, 10 mL; solids available from Aldrich or other chemical supply company)
- Aqueous  $\text{Co}(\text{C}_2\text{H}_3\text{O}_2)_2$  (~0.5 M, 10 mL)
- Stirring rods or paintbrushes, one per solution
- Paper
- Bunsen burner
- Spray bottle of water

### Procedure

1. Use the two different solutions to paint a picture on a piece of regular paper. The picture should appear pink.
2. Hold the paper near a Bunsen burner flame and allow the water to be driven off. The two solutes will appear different colors (blue-green for the chloride, violet for the acetate).
3. If desired, spray the paper with water to demonstrate the return of the pink color upon re-hydration of the cobalt ions.

### Discussion

The following questions can help guide a discussion.

1. Why do the two different salts produce different colors when heated but the same pink color when wet?
2. Based on your answer to question 1, would other metal salts be usable for this demo? What would be necessary in order for the demo to work?
3. Could you reheat and then re-moisten the same painting used in this demo? Explain.

**Hazards**

It is good lab practice to review a chemical's Material Safety Data Sheet (MSDS) before working with any chemical. Follow instructions on the MSDS and encourage students to review them. Wear safety glasses and gloves while preparing solutions and while performing the demonstration. Wash hands thoroughly after performing the demo. Use caution when holding paper near the Bunsen burner flame so as not to burn the paper.

**Disposal**

Check local regulations for proper disposal of chemicals. Cobalt compounds may be recyclable. Check with local hazardous waste removal companies for availability. Heavy metal waste should, whenever possible, be separated from other waste material.

**Halogen Redox Chemistry and Solubility****Objective**

This demonstration takes advantage of the different halogens' colors in their respective elemental forms, and their solubility in nonpolar solvents. Single displacement reactions between organic-soluble halogens and water-soluble halide salts are visible by the color changes of the organic layer. The halogens will displace halides according to their relative reactivities: chlorine will displace both bromide and iodide, leading to brown and purple hexane layers, respectively. Bromine will displace iodide but not chloride; iodine will not displace either.

**List of Materials**

- ~20 mL each of 1.0M aqueous solutions of sodium chloride, sodium bromide and sodium iodide
- ~20 mL hexane
- ~10 mL elemental bromine (or bromine in water)
- ~10 mL chlorine bleach (or chlorine in water)
- ~1 g elemental iodine
- 6 Test tubes or disposable bulb pipets

**Procedure**

1. Have students make observations of all six solutions before mixing.
2. Mix water and hexane to show that it forms two layers instead of mixing.
3. Dissolve each of the halogens in hexane in two test tubes apiece, and have students note the colors of the hexane when each halogen is dissolved in it.
4. To each of the test tubes, add the solution of one of the sodium salts (containing the anion of one of the two other halogens; e.g., add NaBr to one tube with  $\text{Cl}_2$  and NaI to the other).

5. Shake each tube, allow the layers to re-form and have students make note of any color changes.

### Discussion

These questions can help guide students thinking during and after the activity:

1. Is hexane polar or non-polar? How can you tell?
2. For each test tube, explain what is happening.
3. If this demo had included a solution of sodium fluoride, do you think you would have needed to worry about generating fluorine gas? Why or why not?
4. In Unit 4 you learned about electronegativity. Compare the trend in electronegativity with the trend in halogen displacement reactions.

### Hazards

It is good lab practice to review a chemical's Material Safety Data Sheet (MSDS) before working with any chemical. Follow instructions on the MSDS and encourage students to review them. Halogens are toxic, albeit minimally so in the quantities described here. Hexane is flammable. Do not perform this experiment near an open flame. Wear proper protective equipment when performing this demonstration.

### Disposal

Check local regulations for proper disposal. Bromine, iodine, and hexane require hazardous waste disposal according to applicable federal, state, or local regulations. Hexane should be separated from aqueous reagents as much as possible.

## Going Deeper (In-Class Discussion or Reflection)

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Instructors should allow up to 30 minutes for discussion at the end of the session, or students can use the time to reflect on one or more of these questions in journals.

1. Why is it not enough merely to set up two half-reactions with significantly different reduction potentials if you're trying to design a battery?
2. Based on the form of the Nernst equation, why does a reaction coming to equilibrium mean a "dead battery?"
3. Why do you suppose the 18-electron rule is analogous to the octet rule?
4. Look up the structure of the heme group of hemoglobin. Draw the Lewis structures of  $O_2$ ,  $CN^-$ , and  $CO$ , and show how they might be able to bond to the iron in the heme group. How many electrons does this give to the metal?

## Before the Next Unit

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Learners should read the Unit 11 text if they haven't already done so. They may wish to read one or more of the reading assignments from the list below, or, if you choose to have them use the course materials outside of class, they can watch the Unit 12 video and/or read the Unit 12 text as an assignment before the next session.

## References and Additional Resources

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Levi, Primo. *The Periodic Table*. 1st American ed. New York: Schocken Books, 1984.

Dr. Campbell's Favorite Demos. "Vitamin C Cleans Up Iodine Stains on Hand," Accessed August 28, 2013. <http://campbelldemo.blogspot.com/search?q=Vitamin+C+Cleans+Up+Iodine+Stains+on+Hand>

Ford, Leonard Augustine *Chemical Magic*. Revised by E. Winston Grundmeier. 2nd ed. New York: Dover Publications, 1993.

Muske, K. R., C.W. Nigh and R. D. Weinstein, "A Lemon Cell Battery for High-Power Applications." *Journal of Chemical Education* 84: 4 (2007): 635-638.

## For Professional Development

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In addition to watching the videos, reading the text, and going through the activities listed in the course guide, participants taking this course for professional development should read the following papers and answer the corresponding reflection questions. Participants should then complete the accompanying professional development assignments.

### Further Reading & Reflection Questions

Acar, Burcin and Leman Tarhan. "Effect of Cooperative Learning Strategies on Students' Understanding of Concepts in Electrochemistry." *International Journal of Science and Mathematics Education*. 5 (2007), 349-373. Accessed August 13, 2013.

<http://link.springer.com/content/pdf/10.1007%2Fs10763-006-9046-7.pdf>

Chemical Heritage Foundation, "Electrochemistry: Electrolysis of Solutions – Teacher Activity." Accessed October 24, 2014. <http://www.chemheritage.org/Downloads/Classroom-Activities/Teacher-Guide--Electrolysis.pdf>

1. Were you surprised by the results of this study? Why or why not?
2. Do you agree with the author that if cooperative learning is properly applied, as in this study, the formation of students' misconceptions can be prevented? Why or why not?
3. Does this paper influence how you will approach teaching redox chemistry? If so, how? If not, why not?

Barke, Hans-Diester. "Two Ideas of the Redox Reaction: Misconceptions and Their Challenge in Chemistry Education." *African Journal of Chemical Education*. 2:2 (2012), 32-50. Accessed August 26, 2013. <http://www.ajol.info/index.php/ajce/article/view/82445/72600>

1. Have you encountered students holding any of the same misconceptions as those discussed in this paper? Are there some that are more prevalent than others? Were you surprised by any of the misconceptions? If so, which ones and why? If not, why not?
2. Do you agree with the author's suggestions for teaching redox chemistry? If so, which ones, and why? If not, why not?
3. Does this paper influence how you will approach teaching redox chemistry? If so, how? If not, why not?

### Professional Development Assignments

1. After reading the papers above and reflecting on the questions presented, develop a lesson plan designed to teach material presented in this unit.
2. Using a group activity or classroom demonstration presented in this course guide, show how you would implement it in your classroom. Where would it fit into your curriculum or standards? Would you change the demonstration or activity in any way? How would you assess student learning?