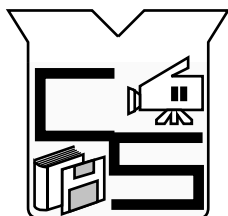


A SourceBook Module

Version 1.0 1994

*Funded in part under
National Science Foundation
Grant No. TPE 88-50632*

*ChemSource Project Principal Investigator:
Mary Virginia Orna, OSU
Department of Chemistry
College of New Rochelle
New Rochelle, NY 10805
Phone: (914) 654-5302
FAX: (914) 654-5387*



ChemSource

*Instructional Resources for Preservice and
Inservice Chemistry Teachers*

OXIDATION-REDUCTION

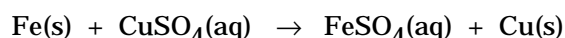
Topic Overview



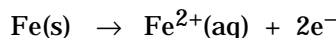
CONTENT IN A NUTSHELL

Oxidation-reduction reactions, or redox reactions, are common. Redox reactions are imagined to involve a *transfer of electrons*. *Oxidation* is always accompanied by *reduction*. Redox reactions are responsible for respiration supporting life, combustion of gasoline, browning of a freshly cut apple, corrosion of metals, operation of an automobile battery, photosynthesis, and many other reactions.

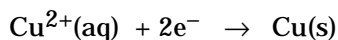
A simple, readily demonstrated example of oxidation-reduction is the reaction of an iron nail with copper(II) ions in a water solution of copper(II) sulfate:



The process can be *imagined* to consist of two steps. The iron atoms lose electrons and become oxidized:

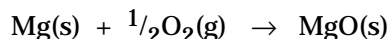


The copper(II) ions gain electrons and become reduced:



The mental picture is one of electron transfer. Oxidation, a loss of electrons, is always counterbalanced by an equal extent of reduction, a gain of electrons.

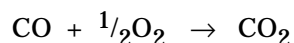
Another example comes from the formation of magnesium oxide, MgO, from the elements magnesium and oxygen.



Think about the model used to explain ionic bond formation. Magnesium loses 2e^{-} to oxygen, thus forming Mg^{2+} ion, while oxygen gains 2e^{-} to become an oxide ion, O^{2-} . We always picture a transfer of electrons in oxidation-reduction reactions. In this case, Mg loses two electrons while oxygen gains two electrons. Losing electrons is oxidation; gaining electrons is reduction.

When an atom either in the elemental state or as an ion loses one or more electrons, it is said to be *oxidized*. Conversely, when the atom gains one or more electrons, it is said to be *reduced*.

Also, when an atom is covalently bonded to other atoms in a molecule or ion, we sometimes describe the resulting partial charge build-up on that particular atom in terms of electron loss (oxidation) or electron gain (reduction). Thus, in a reaction such as



the carbon atom is said to be oxidized because two more of its electrons are shared (unevenly) with the more electronegative oxygen.

The term oxidation once implied chemical reaction with oxygen. Today the term is broadly used to suggest loss of electrons during reaction.

In many cases, it is impossible to say that a particular atom has gained or lost electrons. For this reason, a way to characterize redox reactions is based upon *oxidation numbers*. An oxidation number is a positive or negative number assigned to an atom according to an arbitrary set of rules. Rules to determine oxidation states will be developed in this module.

In the reactions used as examples, the oxidation number of each iron atom is changed from 0 to +2, while the oxidation number of each copper atom is changed from +2 to 0. The oxidation number of each magnesium atom is changed from 0 to +2, while the oxidation number of each oxygen atom is changed from 0 to -2. Using rules given later, the oxidation number of each carbon atom changes from +2 to +4 as CO becomes CO₂.

The definition of reduction is often confusing to students. Originally, reduction referred to obtaining metals (elements) from their ores (compounds). Reduction brings about a decrease in oxidation number.

In a redox reaction, the sum of the increases in oxidation numbers for the oxidized atoms is balanced exactly by the sum of decreases in oxidation numbers for the reduced atoms. This fact can be utilized to balance coefficients in any redox equation. Substances in redox reactions that undergo oxidation are called *reducing agents* while substances that become reduced are called *oxidizing agents*. An oxidizing agent is an *agent* that brings about oxidation and, in so doing, becomes reduced.

This material may follow coverage of periodicity, atomic structure, and conservation of mass and charge. It usually follows stoichiometry, and the first activity requires use of stoichiometry concepts. It should precede any discussion of electrochemistry.

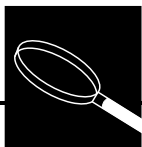
1. Many chemical reactions involve the apparent transfer of electrons between reacting chemical species.
2. Oxidation occurs when a reacting species loses electrons; reduction occurs when a reacting species gains electrons.
3. Oxidation is always accompanied by reduction.
4. The energy released from redox reactions can be much larger than that released from acid base reactions; in that case reactant bonds are relatively weak while product bonds are extremely strong.
5. Some oxidation-reduction reactions occur slowly while others occur very rapidly.
6. Oxidation numbers can be assigned to atoms according to a set of rules.
7. A redox reaction always involves a change in the oxidation number of two or more atoms during a reaction.
8. Oxidation numbers can be used as a "bookkeeping" device in balancing redox equations, predicting further reactivity, and in identifying species that exhibit multiple oxidation states. They are assigned on the basis of associating shared electrons with the more electronegative element.
9. Oxidizing agents are reduced; reducing agents are oxidized.
10. Oxidation-reduction reactions permeate daily life. Examples of common redox reactions range from respiration and combustion to corrosion and to the use of anti-oxidants in foods.

1. Conservation of mass and charge
2. Electronegativity
3. Periodic Table and periodicity of properties
4. Equation balancing by inspection
5. Stoichiometry
6. Bond formation theory

PLACE IN THE CURRICULUM

CENTRAL CONCEPTS

RELATED CONCEPTS



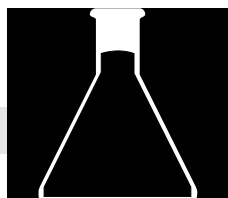
RELATED SKILLS

1. Predict common oxidation numbers using the Periodic Table.
2. Recall charges on common polyatomic ions.
3. Use an algebraic number line.
4. Predict bond formation using electron configurations and Lewis electron dot structures.

PERFORMANCE OBJECTIVES

After completing their study of oxidation-reduction, students should be able to:

1. define using their own words, and with examples, the terms: oxidation, reduction, redox, oxidizing agent, reducing agent, oxidized, reduced.
2. compare the quantity of energy released from the most vigorous redox reactions to that released from the most vigorous acid-base reactions.
3. determine the oxidation numbers of atoms in both reactants and products in simple reactions, identify the oxidizing and reducing agents, and identify the substances oxidized and reduced.
4. use the concept of oxidation number together with the concept of a number line to illustrate how oxidation and reduction always occur to the same extent.
5. cite examples in which redox reactions are involved in biochemistry, production of common goods, and environmental changes.
6. balance simple oxidation-reduction equations by at least one method (oxidation number or half-reaction method).



Concept/Skills Development

Activity 1: Oxidation States of Manganese

(Modified from Experiment C15 of Doing Chemistry)

Introduction

Atoms in compounds may be found in various oxidation states. Atoms of some elements, like manganese, have many possible oxidation states. Oxidation-reduction reactions lead to changes in oxidation states.

Purpose

The oxidation state of manganese in potassium permanganate is changed through a series of redox reactions. The redox reactions are controlled by adjusting the acidity of the reaction medium.

Safety

1. Wear protective goggles throughout the laboratory activity.
2. Dilute sulfuric acid and dilute sodium hydroxide solutions are corrosive. Solid potassium permanganate is a strong oxidant. Potassium permanganate and sodium bisulfite are toxic.
3. Use caution when handling the dilute sulfuric acid solution.
4. Do not permit the potassium permanganate to come in contact with any reducing agents. Save any spilled material for disposal according to the procedure given.
5. Dispose of all materials as your teacher directs.

Procedure

Use pulled Beral pipets

1. Put 10 drops KMnO_4 solution in four wells of a 24-well plate, A-1 through A-4.
2. Add four drops NaOH solution to well A-2.
3. Add three drops 3M H_2SO_4 solution to well A-4.
4. Place the plate on a sheet of white paper or other white background.
5. Add three drops NaHSO_3 solution to well A-2 while stirring with a clean toothpick until there is evidence of a reaction.
6. Add three drops NaHSO_3 solution to well A-3 while stirring with a toothpick. Note evidence for a reaction.
7. Slowly add NaHSO_3 solution to well A-4, while stirring with a toothpick, until there is evidence of a reaction.
8. MnO_2 is an insoluble brown solid. Mn^{2+} is faintly pink in solution. MnO_4^{2-} is dark green in solution. Based upon these colors, identify the manganese reaction products formed in wells A-2, A-3, and A-4.
9. Thoroughly wash your hands before leaving the laboratory.

LABORATORY ACTIVITY: STUDENT VERSION



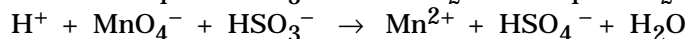
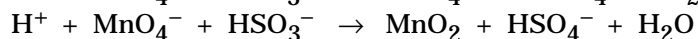
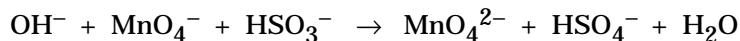


Data Analysis and Concept Development

1. For each well, determine whether the Mn in MnO_4^- undergoes oxidation or reduction.
2. For each well, decide whether MnO_4^- acts as an oxidizing agent or a reducing agent.
3. Identify the other agent in the reaction, and indicate whether it is an oxidizing agent or a reducing agent.
4. Predict the effect of replacing NaHSO_3 with NaHSO_4 . Test your prediction.
5. Complete the table involving reactions between KMnO_4 and NaHSO_3 .

	Well	Change in Oxidation State of Mn	Change in Oxidation State of S	Electrons Gained by Mn	Electrons Lost by S
Reaction 1	A-2	7 \rightarrow 6	4 \rightarrow 6	1	2
Reaction 2					
Reaction 3					

6. In Reaction 1 of the table above, each MnO_4^- takes up one electron, but each HSO_3^- gives up two electrons. Account for this difference.
7. Predict the stoichiometric ratio of MnO_4^- to HSO_3^- in Reactions 2 and 3.
8. Balance each of the following skeleton reactions based upon the stoichiometric ratios developed in Question 7:



Implications and Applications

This activity involves observing a reaction in which oxidation states change. Color changes provide concrete evidence for reaction. The observed ratios suggest ways in which one can keep track of electrons gained or lost, and serve as the basis of a procedure for balancing oxidation-reduction equations.

Activity 1: Oxidation States of Manganese**Major Chemical Concept**

The oxidation state of manganese in potassium permanganate is changed as the result of oxidation-reduction reactions.

Level

Introductory level for basic, general and honors students.

Expected Student Background

See *Related Concepts*. Students should have experience with small-scale equipment (8- or 12-well strips or 96-well plates; obtaining drops from Beral storage devices. See Frame A36950 and subsequent frames of the *Doing Chemistry A-side* videodisc for demonstrations of Beral pipet preparation.)

Time

One 50-minute class period.

Safety

Goggles and aprons worn. Wash spills of H_2SO_4 and NaOH immediately. Do not ingest chemicals. Dispose of manganese compounds by converting to Mn^{2+} and flushing down the drain with large amounts of water or converting to MnO_2 and disposing in accord with local regulations. Although not a part of the activity, teachers should know that mixtures of potassium permanganate and concentrated sulfuric acid lead to the formation of explosive manganese oxide, Mn_2O_7 .

Materials (For 24 students working in pairs)

- 0.01 M KMnO_4 , potassium permanganate, 250 mL (0.4 g KMnO_4 per 250 mL solution)
- 3 M H_2SO_4 , sulfuric acid, 250 mL (42 mL conc. H_2SO_4 per 250 mL solution—add acid to water)
- 2.0 M NaOH , sodium hydroxide, 250 mL (20 g NaOH per 250 mL solution)
- 0.01 M NaHSO_3 , sodium bisulfite, 250 mL (0.26 g NaHSO_3 per 250 mL solution)
- 0.01 M NaHSO_4 , sodium bisulfate, 250 mL (0.30 g NaHSO_4 per 250 mL solution)
- Twelve 24-well plates (or small test-tubes)
- Beral pipets (or droppers)
- White (bond) paper
- Round toothpicks

Advance Preparation

Prepare the 0.01 M KMnO_4 and NaHSO_3 fresh. Within 48 hours of the laboratory activity, transfer these solutions to Beral storage pipets. (The KMnO_4 and NaHSO_3 solutions are reactive and change upon storage.)

Pre-Laboratory Discussion

Suggest that atoms of one element may have different oxidation states depending upon the compound. Stress the importance of making detailed observations, noting colors, textures, bubbles, *etc.* The number of drops of NaHSO_3 may need to be increased to obtain the proper colors.

**LABORATORY
ACTIVITY:
TEACHER
NOTES**



Question 4 of the Data Analysis Section asks students to predict the effect of replacing NaHSO_3 with NaHSO_4 . The easiest way to test this prediction is to repeat the procedure replacing NaHSO_3 with NaHSO_4 . If students confuse these two substances, they will get unexplainable results. Therefore, during pre-laboratory discussion, you may want to point out this potential misidentification problem.

Teacher-Student Interaction

No special monitoring is required. Help students focus upon detailed observations and recording notes as you walk around the laboratory.

Anticipated Student Results

The alkaline solution turns green and then slowly turns murky. A brown precipitate forms in the well with no added acid or base. The acid solution turns colorless.

Answers to Data Analysis and Concept Development

1. MnO_4^- is reduced in each well.
2. MnO_4^- is the oxidizing agent in each well.
3. HSO_3^- is the reducing agent in each well and is oxidized.
4. In HSO_4^- , the sulfur is already oxidized to the +6 oxidation state. This ion is not a reducing agent.
- 5.

	Well	Change in Oxidation State of Mn	Change in Oxidation State of S	Electrons Gained by Mn	Electrons Lost by S
Reaction 1	A-2	7 → 6	4 → 6	1	2
Reaction 2	A-3	7 → 4	4 → 6	3	2
Reaction 3	A-4	7 → 2	4 → 6	5	2

6. Two permanganate ions react for each bisulfite ion that reacts.
$$2 \text{OH}^-(\text{aq}) + 2 \text{MnO}_4^-(\text{aq}) + \text{HSO}_3^-(\text{aq}) \rightarrow 2 \text{MnO}_4^{2-}(\text{aq}) + \text{HSO}_4^-(\text{aq}) + \text{H}_2\text{O}(\text{l})$$

This may be thought of as each manganese atom lowering its oxidation number from +7 to +6 as the result of picking up one electron, while, at the same time one sulfur atom releases 2 electrons thereby raising its oxidation number from +4 to +6.
7. The ratios are $2 \text{MnO}_4^- / 1 \text{HSO}_3^-$ in A-2; $2 \text{MnO}_4^- / 3 \text{HSO}_3^-$ in A-3; and $2 \text{MnO}_4^- / 5 \text{HSO}_3^-$ in A-4.
8. A-2: $2 \text{OH}^-(\text{aq}) + 2 \text{MnO}_4^-(\text{aq}) + \text{HSO}_3^-(\text{aq}) \rightarrow 2 \text{MnO}_4^{2-}(\text{aq}) + \text{HSO}_4^-(\text{aq}) + \text{H}_2\text{O}(\text{l})$
A-3: $2 \text{H}^+(\text{aq}) + 2 \text{MnO}_4^-(\text{aq}) + 3 \text{HSO}_3^-(\text{aq}) \rightarrow 2 \text{MnO}_2(\text{s}) + 3 \text{HSO}_4^-(\text{aq}) + \text{H}_2\text{O}(\text{l})$
A-4: $6 \text{H}^+(\text{aq}) + 2 \text{MnO}_4^-(\text{aq}) + 5 \text{HSO}_3^-(\text{aq}) \rightarrow 2 \text{Mn}^{2+}(\text{aq}) + 5 \text{HSO}_4^-(\text{aq}) + 3 \text{H}_2\text{O}(\text{l})$

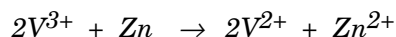
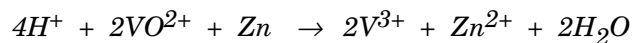
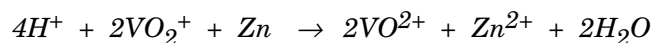
Post-Laboratory Activities

Discuss the observations, working toward a consistent set of class observations. Discuss the assignment of oxidation states.

Assessing Laboratory Learning

Ask specific questions about the substances and reactions studied. Also pose open-ended items such as:

1. NaHSO_3 is frequently used to remove brown rust stains. Suggest a reason for this use. What must you know about the rust stains to make this suggestion consistent with principles? *Iron(III) oxide and hydroxide are much less soluble in acid solution than iron(II) hydroxide and oxide. The NaHSO_3 reduces iron(III) to iron(II).*
2. Before safety concerns (high toxicity) caused it to be dropped from common use, laboratory activities were performed with ammonium vanadate, $(\text{NH}_4)_3\text{VO}_4$. Acidic yellow solutions of this compound were shaken with solid zinc metal, Zn. The solutions underwent spectacular color changes. Suggest reasons that might account for these changes, and describe the kind of information you would need to have or experiments you would want to perform to confirm your hypotheses. *Zinc metal reduces vanadium(V) through vanadium(IV) and vanadium(III) intermediates to vanadium(II).*



Each species has a unique, characteristic color in acid aqueous solution. One way to confirm this series of reactions would be to measure the potential of the solution during the reduction process and relate potential changes to color changes. The reverse titration is often performed. Solutions of vanadium(II) are titrated with cerium(IV) to produce, in stepwise fashion, vanadium(III), vanadium(IV), and finally vanadium(V).



LABORATORY ACTIVITY: STUDENT VERSION

Activity 2: The Silver Mirror Reaction

Reference: The Woodrow Wilson microscale version of this lesson was developed by Fen Lewis, Strongsville High School, 7701 Beach Road, Wadsworth, OH 44281, and was based upon Kemp and Marwin. (1981). Journal of Chemical Education, 58, 655 .



Introduction

Everyday mirrors that we take for granted are manufactured using redox reactions in which silver compounds are reduced to silver metal. Because the silver metal becomes bonded to the glass so that tarnishing of the silver is slow, a shiny 'silver mirror' lasts for months or years, depending on methods of preservation.

Purpose

An ordinary silver mirror is produced using typical laboratory materials.

Safety

1. Wear protective goggles throughout the laboratory activity.
2. Nitric acid is corrosive. Silver nitrate causes stains.
3. Handle these substances with caution. Wash spills immediately with large amounts of water. Hold your thumb on the stopper while shaking.
4. Always mix the solutions fresh and dispose of them immediately after use with large amounts of water. The materials may form explosive silver fulminate, $\text{Ag}_2\text{C}_2\text{N}_2\text{O}_2$, on standing. *Never* premix the reagents.
5. Dispose of all materials as your teacher directs.

Procedure

1. Thoroughly clean with laboratory detergent the inside of a test-tube given to you for this laboratory activity. A scrupulously clean test-tube is essential for successful completion of this activity.
2. Rinse the test-tube with distilled water. Taking care to keep the acid off your skin, rinse again with 1/4 test-tube full of 6 M nitric acid, followed by a rinse with distilled water. Discard under hood.
3. Rinse again with acetone, then with distilled water.
4. Add the following amounts of solutions in the *exact* order listed:
16 Drops stabilized honey solution (5%). Roll the tube to wet it with this solution before adding silver nitrate.
8 Drops 8.0% silver nitrate solution
8 Drops 12% ammonium nitrate solution
16 Drops 10% sodium hydroxide solution
5. Quickly stopper the test-tube with a cork and shake it while holding the cork in place. The inside surface of the test-tube should be wetted for a good coating. Continue shaking the test-tube for about 3 min. Observe the changes in appearance for 5 min.
6. Wash the solution down the drain with 1 L water. This is an important safety precaution to prevent the possible formation of an explosive mixture (after standing many hours or days). Rinse the mirrored test-tube gently but thoroughly with distilled water. Allow the tube to air dry.
7. Thoroughly wash your hands before leaving the laboratory.

Data Analysis and Concept Development

1. What is the visible product of the reaction?
2. Did the silver ions gain or lose electrons in the reduction process?
3. An oxidation reaction must take place along with a reduction reaction. Honey (actually the sugars in honey) is the partner for the reaction of silver ions in this activity. What happened to the sugar molecules?

Implications and Applications

The production of mirrors is the direct application illustrated by this activity. Essentially all metallurgical processes involve redox reactions. When metals are obtained from ores or are purified, the process involves reducing the element to its metallic state or oxidizing some impurity to a compound that can be separated from the remaining metal.



LABORATORY
ACTIVITY:
TEACHER
NOTES

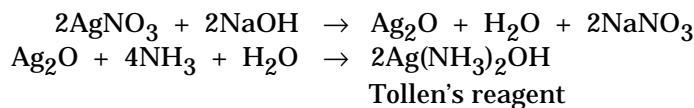
Activity 2: The Silver Mirror Reaction

Major Chemical Concept

The controlled reduction (*i.e.*, slow reduction) of silver ion leads to the formation of metallic silver.

Reactions:

Formation of the reagent (Tollen's Reagent):



Reduction of the reagent:



Level

Introductory level for basic, general, and honors students.

Expected Student Background

See *Related Concepts and Skills*.

Time

One 50-min class period

Safety

1. Always mix the chemicals fresh and dispose of them immediately after use with large amounts of water. The chemicals may form explosive silver fulminate, $\text{Ag}_2\text{C}_2\text{N}_2\text{O}_2$, on standing. *Never* premix the chemicals.
2. Sometimes teachers use acetone to clean glassware thereby removing grease-containing materials from the glass surface. Also, some teachers recover used glassware by removing silver with concentrated nitric acid. Mixtures of these two materials, acetone and concentrated nitric acid, are explosive.

Materials (For 24 students working in pairs)

24 Test-tubes, 15- x 120-mm, and corks to fit them.

5% Honey, 200 mL (10 g honey in 100 mL water. Add 1.2 g tartaric acid, boil, then cool the solution. Add 20 mL ethanol and dilute with water to 200 mL)

8% Silver nitrate, AgNO_3 , 100 mL (8.0 g AgNO_3 per 100 mL solution)

12% Ammonium nitrate, NH_4NO_3 , 100 mL (12.0 g NH_4NO_3 per 100 mL solution)

10% Sodium hydroxide, NaOH , 200 mL (20 g NaOH per 200 mL solution)

6 M Nitric acid, HNO_3 (add 76 mL concentrated nitric acid to 80 mL distilled water and dilute to 200 mL)

Acetone, 200 mL

Advance Preparation

Obtain the honey from a food store.

Pre-Laboratory Discussion

Focus the pre-laboratory discussion on practical applications of redox reactions and controlling the rate of a reaction. Combustion reactions, for example, are usually very fast redox reactions while rusting and tarnishing are slow.

Teacher-Student Interaction

Pay particular attention to safety. Several dangerous substances are in use. As you check individual students, check that the test-tubes have been properly cleaned. Dirty test-tubes are a main cause for failure and/or poor results. You may also need to coach the shaking process to make certain that adequate shaking occurs without making it too vigorous. (You may ask students to place corked tubes inside plastic zip-closure bags for shaking. Have them hold the cork in place while shaking even if a zip-closure bag is used. This strategy is helpful when the material being shaken presents either toxicity or stain-producing hazards.)

Anticipated Student Results

A laboratory full of beautiful silver mirrors will appear after about 35 min of work.

Answers to Data Analysis and Concept Development

1. Metallic silver
2. Silver ions gained electrons in reduction
3. Sugar molecules are oxidized.

Post-Laboratory Discussion

Discuss the original definition of the term reduction (reducing a metal to an element from a compound or ore) in the context of mirror formation. Point out that there are organic and biochemical reducing agents such as glucose, as well as inorganic reagents such as NaHSO_3 , and that the energy for living systems is derived from biochemical redox reactions. Discuss the relative quantity of energy released from this reaction as compared with, say, the energy released in combustion of methane in a burner.

Assessing Laboratory Learning

Ask specific questions about the chemicals and reactions. Consider asking a safety question about the storage of Tollen's reagent.

Demonstration 1: Reduction of Copper(II) Oxide

(Modified from *Experiment A45 of Doing Chemistry*)

Description

Black copper(II) oxide is reduced to form copper metal (with its characteristic appearance) using methane gas. The reaction is clean and stoichiometric.

Materials

Burner
 Hose—with bent glass tubing inserted into rubber stopper or cork for clamping
 Test-tube Pyrex™, 25- x 200-mm (16- x 100-mm)
 Wire to suspend tube for weighing
 Copper(II) oxide, finely powdered, 1 g
 Source of methane gas
 2 Single-buret clamp
 2 Support stands
 Balance
 May need Y-tubes and clamps to provide sufficient gas outlets
 Safety shield

DEMONSTRATIONS



Laboratory Hints

An interesting variation is to use a tightly fitting two-holed rubber stopper for the tube. Insert a long glass tube through one hole so that it comes within 1 cm of the bottom of the test-tube and extends 5 cm to the other side of the stopper. Insert a 10-cm length into the other hole so that 3.5 cm of tubing protrudes on each side of the stopper. Connect rubber tubing from the gas source to the long glass tube. Connect rubber tubing from the burner to the short glass tube. Turn on the gas. The gas first flows over the oxide, and then flows out through the burner. Wait 15 seconds. Light the gas. When the reaction is complete, remove the burner, and allow the reaction test-tube to cool while the gas burns in the burner. Y-tubes or T-tubes and suitable clamps may be needed to provide enough gas outlets in laboratories where these are not available.

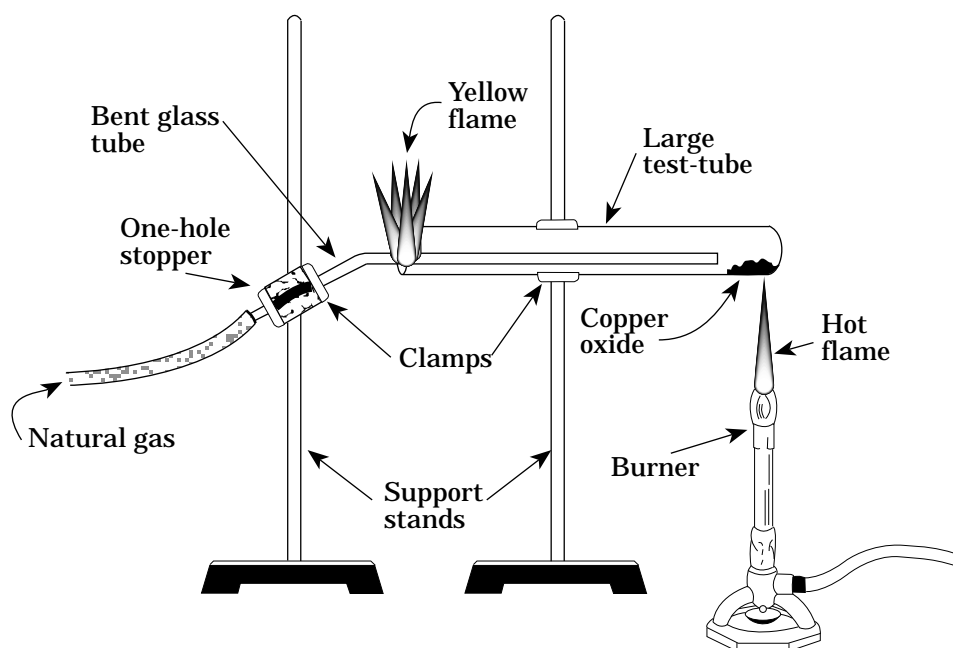


Figure 1. Apparatus for the reduction of copper(II) oxide.

Discussion

A possible error source is partial reduction of black CuO to red Cu_2O , which is often mistaken for Cu .

Safety

Natural gas is flammable and potentially explosive. Hot glassware can cause burns. Bending glassware and inserting glass tubing through rubber stoppers can lead to severe cuts.

Precautions

1. Use a safety shield.
2. Ignite escaping gas after a 15-second wait. Be careful handling hot glassware; use fingers to sense whether glassware has cooled to room temperature before grasping. Do not use plastic coated clamps.
3. To insert tubes into stoppers, use fire-polished glass (*i.e.*, glass heated to remove sharp edges.) Use a lubricant such as glycerin on the glass and the stopper. Wrap the glass and the stopper in a cloth towel. Slowly, with minimal pressure, twist the glass into the stopper.

Disposal

Reuse the same test-tube from year to year. The solid may be disposed of with ordinary solid waste.

Procedure

1. Set up the apparatus behind a safety shield.
2. Prepare a wire from which to hang a test-tube. Wrap a loop of wire around the top of a 16-mm test-tube near the lip. Shape the other end of the wire into a loop. This loop should be about 3-4 cm from the tube. Determine the mass of a test-tube by hanging it from the loop of wire attached to the balance hook.
3. Weigh a clean, dry test-tube. Record the mass.
4. Add about 1 g copper(II) oxide with a metal scoop so that as little as possible of the powder gets on the side of the test-tube.
5. Reweigh the test-tube and powder. Record the mass.
6. Insert the glass tube connected to the methane gas supply carefully so that it does not touch the powder. You may want to hold it during the demonstration or clamp it to a support stand. The glass tube should extend far into the test-tube, but not touch either the solid in the test-tube or the test-tube itself.
7. Turn on the source of methane gas for the glass tube inserted into the test-tube. After five seconds ignite the gas at the mouth of the tube. Adjust the methane flame at the mouth of the tube so that it is about 5 cm (2 in) high.
8. Light a burner. Heat the exterior of the test-tube to as high a temperature as possible with the burner for about 5 to 7 min or until copper color appears in the material in the test-tube.
9. Turn off the burner, but keep the methane flowing through the inside of the test-tube until the test-tube cools down (about 5 min). Once cooled, turn off the methane. (Because of the heat released at the top of the tube when the methane is burning, the tube will not cool to room temperature. Once the gas is turned off, wait until the tube cools to room temperature.)
10. Reweigh the test-tube and material remaining. Record the mass.

Sample Data Table

Mass empty tube	41.80 g
Mass tube plus oxide	43.00 g
Mass tube after reduction	42.78 g

Data Analysis

The mass of copper(II) oxide is: $43.00\text{ g} - 41.80\text{ g} = 1.20\text{ g}$

The mass of solid remaining is: $42.78\text{ g} - 41.80\text{ g} = 0.98\text{ g}$

If all remaining material were Cu, there would be:

$$0.98\text{ g Cu} \times (1\text{ mol Cu} / 63.54\text{ g Cu}) = 0.0154\text{ mol Cu}$$

If all the lost mass were due to oxygen, there would be:

$$1.20\text{ g} - 0.98\text{ g} = 0.22\text{ g O, and}$$

$$0.22\text{ g O} \times (1\text{ mol O} / 16.00\text{ g O}) = 0.0138\text{ mol O}$$

The mole ratio of Cu to O is:

$$0.0154\text{ mol Cu} / 0.0138\text{ mol O} = 1.16$$



Questions

1. Write a balanced chemical equation to describe the reaction between methane and copper(II) oxide. $[CH_4 + 3CuO \rightarrow CO + 2H_2O + 3Cu]$
2. Use the sample data here. Assume that the formula of copper(II) oxide is CuO. For the sample size taken, predict the mass of copper that should remain after reaction.

$$[1 \text{ mol Cu} \times (63.54 \text{ g Cu} / 1 \text{ mol Cu}) = 63.54 \text{ g Cu}$$

$$1 \text{ mol O} \times (16.00 \text{ g O} / 1 \text{ mol O}) = 16.00 \text{ g O}$$

$$63.54 \text{ g Cu} + 16.00 \text{ g O} = 79.54 \text{ g CuO}$$

$$1.20 \text{ g CuO} \times (63.54 \text{ g Cu} / 79.54 \text{ g CuO}) = 0.959 \text{ g Cu}$$

For CuO, the predicted mass of Cu remaining is 0.959 g.]

3. Compare the predicted mass of copper remaining with that actually observed. Suggest possible reasons for discrepancies. [The observed mass is 0.98 g Cu, and the predicted mass is 0.96 g Cu. This difference can be due to weighing error or insufficient heating.]

Demonstration 2: Sodium in Water

(Modified from Experiment C03 of Doing Chemistry)

Description

A small piece of metallic sodium is placed in a dish of water containing phenolphthalein on an overhead projector stage. The sodium reacts vigorously, melts, and sputters on the surface of the water as a sphere of continuously decreasing size until it has completely reacted. The moving sodium leaves a pink trail in the water.

Materials

Overhead projector and screen
Shield
Petri dish (or 1-L beaker)
Stainless metal spatula with small blade
Sodium metal (lump the size of a small, dried, split, green pea)
Water (ice)
1% Phenolphthalein (in 95% ethanol)
Dry paper towels
Distilled water
Tongs or forceps

Safety

Metallic sodium is very dangerous; it is a strong reducing agent and prone to highly exothermic redox reactions. Fires and burns are possible. Flying debris from explosions resulting from misuse of sodium may cause severe injury.

Precautions

1. Store the sodium in small amounts under kerosene or mineral oil (nonproton-donating, nonreactive solvents). Store only enough to use for this demonstration for the school year, and no more than 10 g. Wrap the bottle in a heavy-weight polyethylene bag. Store in a place where it will not get wet, or come in contact with acids or oxidizing agents. Check every three months to be sure that the storage procedures are being followed.
2. Do not touch sodium with your fingers.

3. Use only the amount recommended. The amount used is magnified by the overhead projector. Your students will almost certainly try to tempt you use a larger piece of sodium; *never* do this. Wear goggles; have all students wear goggles. Use cool water.

Disposal

1. Call for assistance (such as district science supervisor, local college chemistry department, local fire marshal) in destroying excess sodium. Sodium stored in small amounts according to the precautions noted presents no special danger or hazard to the school, students, custodians, or fire fighters.
2. After the reaction ceases, the reaction mixture may be disposed of safely in the sink.
3. Be certain that no small metallic sodium fragments remain on the towel before disposing of it. If fragments are noticed, use tongs to immerse the towel in a large plastic bucket filled with cold water.

Procedure

1. Set up an overhead projector. Place a safety shield between the projector and the students. Work from the student side of the projector to protect yourself. Wear goggles and have your students wear goggles. Place a Petri dish (or a 1-L beaker) on the projector stage. Pour some cool water into the dish, and add 1-2 drops of 1% phenolphthalein in ethanol. Cool with ice if necessary. Check for visibility by students, and adjust as needed.
2. Remove the bottle of sodium metal from its storage area. Place on a safe, secure surface. Unscrew the cap. Use a dry spatula to “stab” a small chunk of the metal. Lift this chunk out of the bottle with tongs or forceps and place on a dry paper towel.
3. Use the spatula to cut a very small piece of sodium from the chunk. The piece should have a volume no larger than a rice grain. Note the surface of the freshly cut sodium. Bring this to your students’ attention. Have them come close to view the surface, but under no circumstance let them touch the sodium. Do not pass the sodium around the room.
4. Return the chunk of metal to the storage oil as soon as a small piece has been cut off. Tighten the cap.
5. Place the small piece of sodium on the tip of the spatula. Transfer to the Petri dish. Note evidence for reaction.
6. Wear eye protection while disposing of the contents of the Petri dish at the sink. Flush with moderate amounts of water.

Demonstration 3: The Blue Bottle Reaction

(Modified from Experiment A06 of Doing Chemistry)

Description

One solution is divided between two containers (flasks) such that one container is nearly full and the other nearly empty. Shaking the filled container does not seem to produce a dramatic change, but shaking the partially filled container does. Pouring the liquids from the filled container to the empty one and repeating the procedure leads to the same result—only the nearly empty container gives a dramatic change.



Materials

2 Florence flasks, 1-L
Beaker, 2-L
Stirring rod, 40-cm
Glucose (or fructose), 10 g
Sodium hydroxide, NaOH, 10 g
1% methylene blue, in dropping bottle (dissolve 1 g methylene blue in 100 mL distilled water)
Distilled water

Laboratory Hints

1. Sucrose (table sugar) is not a suitable substitute for glucose.
2. Make a fresh glucose solution on the day of the laboratory activity.
3. This laboratory activity can be performed by students. Students may shake the solutions in stoppered test-tubes. Emphasize the safety precautions. (Remember, every student in the class will be shaking a strongly alkaline solution in a glass container!)

Safety

The solution is alkaline (about 0.2 M NaOH) and a skin irritant with prolonged exposure. Methylene blue will stain clothing.

Precautions

Wear eye protection. Wash off spills. Wear old clothing or laboratory aprons. Check that the safety shower and eye wash fountain are in working condition.

Disposal

Neutralize the sodium hydroxide solution with vinegar; dispose of at the sink.

Procedure

1. Place 1500 mL distilled water in a 2-L beaker. Add to the water 10 g glucose, 10 g sodium hydroxide, and sufficient drops of 1% methylene blue to color the mixture blue.
2. Fill a 1-L Florence flask with this solution. Pour the remaining solution into a second 1-L Florence flask.
3. Swirl each flask vigorously, and note results.
4. Pour from the filled flask into the partially filled flask and swirl the contents of both flasks vigorously. Note observations.
5. Pour once again, shake once again, and note results.

Discussion

Ask students to account for their observations. Once discussion slows down, assuming no one picks up on the role of air, suggest that one of the reactants, which they did not consciously add to the system, might be invisible. Ask, "What reactant should be added to the list?" Generate suggestions for testing their hypotheses.

Background

1. Oxygen oxidizes glucose well in a strongly alkaline solution. Methylene blue is also subject to oxidation. The oxidized form is blue; the reduced form (leuco methylene blue) is colorless. When oxygen is present, methylene blue is blue. When the solution is permitted to stand without shaking, the glucose reduces the methylene blue to the colorless form.

2. Shaking the partially filled flask causes oxygen gas to dissolve and to oxidize the methylene blue. When there is no source of fresh air—a filled flask, or one with very little surface—the blue does not form. You will probably be able to see a bluish color at the top surface of the solution. This careful observation provides support for the mechanism suggested to account for these color changes. In a sealed flask, once all of the oxygen gas is used up, there is no longer a changing of the color. Care must be taken to use fresh solutions.
3. Readily oxidized sugars may replace the glucose. Several redox dyes may replace methylene blue.
4. This demonstration can be used as an effective model for lakes in which there has been eutrophication (enrichment). When the environment is very rich, algae blooms and excessive weeds lead to oxidation of organic matter and subsequent oxygen depletion.

Small-scale

Some teachers implement this as an activity in which students operate at the test-tube scale. Students must wear eye protection at all times during the microscale activity. One way to present the laboratory activity is to provide one filled and one half-filled 4-mL screw cap vial and ask whether “the two solutions are the same.”

Demonstration 4: Hydrogen Peroxide as an Oxidizing and Reducing Agent

Materials

2% Potassium iodide, KI (2 g KI per 100 mL solution)
 0.02 M Potassium permanganate, KMnO_4 (0.3 g KMnO_4 per 100 mL solution)
 3% Hydrogen peroxide, H_2O_2 (commercially available)
 3 M Sulfuric acid, H_2SO_4 (Add 1 mL concentrated H_2SO_4 slowly to 5 mL distilled water.)

Safety

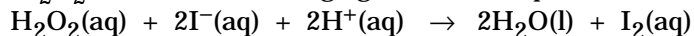
All of the substances are toxic. All except KI are corrosive.

Procedure

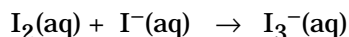
1. Place 100 mL KI solution in a large beaker. Add 6 drops 3 M H_2SO_4 . Add 3% H_2O_2 until a brown color is produced.
2. Place 100 mL KMnO_4 solution in a large beaker. Add 6 drops 3 M H_2SO_4 . Add 3% H_2O_2 until the purple color disappears.
3. This procedure works very well on an overhead projector.

Background

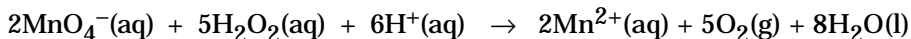
When H_2O_2 acts as an oxidizing agent, water is produced.



The iodine reacts with I^- to produce a colored complex, I_3^- .



When H_2O_2 acts as a reducing agent, oxygen is produced:

*Disposal*

Treat the Mn^{2+} solutions with KMnO_4 in neutral solution. Add sodium bisulfite. Filter the solid MnO_2 and hold for disposal with other solid heavy metal wastes.



Demonstration 5: Displacement of Silver by Copper

(Modified from *Experiment A44 of Doing Chemistry*)

Description

Solid silver metal can be obtained from a reaction in which metallic copper is the reducing agent and is oxidized. Copper is usually considered an easily reduced element. This chemical contrast together with the very attractive reaction (beautiful silver crystals, blue colored solution) make for an excellent demonstration.

Materials

0.1 M Silver nitrate, AgNO₃ (1.7 g per 100 mL solution)
Copper wire, 30 cm,
Graduated cylinder, 100-mL

Safety

Silver nitrate is toxic and causes dark stains.

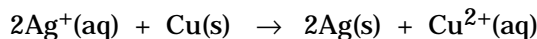
Precautions

Handle silver nitrate solutions cautiously. Wipe up spills immediately.

Procedure

Place 100 mL 0.1 M silver nitrate in a 100-mL graduated cylinder. Coil the copper wire into a spring-like shape with a coil diameter smaller than the internal cylinder diameter. Place the copper coil in the silver nitrate solution. Note evidence for reaction. Allow to stand for several hours.

Reaction



Demonstration 6: Potassium Permanganate and Glycerol

(*Chemical Demonstrations*, B. Z. Shakhshiri, Volume 1, p. 83; shown on Redox videodisc, A27)

Description

Solid potassium permanganate reacts with certain alcohols. The reaction with glycerol is dramatic because it starts slowly. As the reaction proceeds and the mixture heats up, it becomes extremely vigorous. The reaction rate depends upon the size of the potassium permanganate particles. Finely ground potassium permanganate reacts more quickly than does a sample with large crystals. In a side-by-side comparison, placing similar amounts of glycerol on two separate piles of potassium permanganate with different crystal sizes, the effect of surface area on reaction rate can be demonstrated.

Materials

Potassium permanganate, KMnO₄, 5 g
Glycerol, 2 mL
Dropping device (Beral pipet)
Evaporating dish, 90-mm
Spatula
(For reaction rate activity, need mortar and pestle and two of each listed item.)

Safety

Potassium permanganate is toxic and causes burns and stains. The reaction is extremely exothermic and can cause burns and uncontrolled fires.

Precautions

Handle the potassium permanganate cautiously. Perform the reaction away from combustible materials. Keep a loaded fire extinguisher nearby.

Procedure

Place 5 g KMnO_4 in an evaporating dish. Use a spatula to make a small indentation in the pile of solid KMnO_4 . Add 1 mL glycerol to the indentation in the center of the pile of solid. Observe for several minutes. (NOTE: Varying particle size of KMnO_4 provides an interesting lesson on surface area vs. reactivity.)

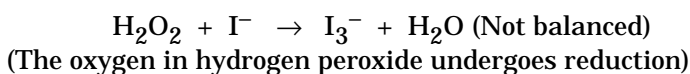
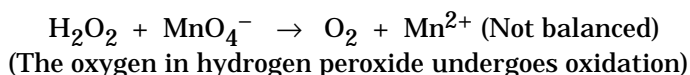
Reactions

Several reactions are possible, such as:



Key Questions

1. Can an atom be oxidized in the absence of oxygen? [Yes. Chemically speaking, oxidation is not linked to oxygen. Oxidation is a loss of electrons.]
2. Why aren't oxidizing agents oxidized? [Oxidizing agents, the agents that bring about oxidation, bring about a loss of electrons by taking in electrons. They are reduced because they act as oxidizing agents.]
3. Why aren't reducing agents reduced? [Reducing agents bring about reduction by giving up electrons. To give up electrons is to become oxidized.]
4. Using a number line, explain why gaining electrons does not make the resulting particle more positive. [Gaining electrons moves to the left on the number line reducing the oxidation number. When the number line concept is used, gaining electrons is treated algebraically—gaining something negative.]
5. How can an atom gain electrons and be “reduced?” [This is a quirk in terms. In everyday usage, “reduction” implies a loss. By defining reduction as a gain in electrons, chemists create confusion with the everyday use of the word reduction.]
6. Explain how it is possible for a substance to act as an oxidizing agent in one case and a reducing agent in another case. [If a substance gives electrons in one reaction but takes electrons in another, it is a reducing agent in the first case, and an oxidizing agent in the second.]



7. Explain why all reactions cannot be classified as redox reactions. [Some reactions are acid base reactions. Other reactions involve forming solids by ‘replacements.’ These reactions do not involve changes in oxidation numbers of atoms.]
8. How can a redox reaction be identified? [A redox reaction is always identified by identifying a change in oxidation number.]
9. How can one infer that bodily exercise increases oxidation of stored food? [Exercise leads to weight loss. The rate of respiration during exercise increases, so that oxygen consumption increases and carbon dioxide production increases.]

**GROUP AND
DISCUSSION
ACTIVITIES**



Counterintuitive Examples

Magnesium Burning in Carbon Dioxide (shown on Redox videodisc, A08)

Description

Magnesium is a strong enough reducing agent to reduce CO_2 to carbon. Carbon dioxide is regarded as something that extinguishes fires, but here supports the burning of magnesium. When something burns in an atmosphere of carbon dioxide, it is a counterintuitive event.

Materials

Dry ice (or sodium bicarbonate, marble chips, and vinegar)
Beaker, large
Wood splint
Magnesium, Mg, ribbon
Tongs
Burner and ignition source

Hazards

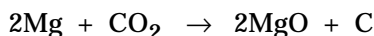
Burning magnesium causes burns. Magnesium ribbon can cause cuts. Dry ice can cause burns. Effervescing acid/carbonate mixtures can spray corrosive liquids.

Precautions

Wear goggles. Handle burning magnesium with tongs. Do not touch burning materials.

Procedure

1. Place dry ice in a beaker and wait about 10 min. (Alternatively, place about 30 g sodium bicarbonate and 15 g marble chips in a large beaker and carefully add 200 mL vinegar.) Ignite a wood splint in a burner flame, and plunge the burning splint into the gas at the top of the beaker. Ignite a piece of magnesium ribbon in a burner flame, and repeat.
2. The wood splint stops burning. The magnesium ribbon continues to burn, and a black solid (carbon) is produced.



In this reaction, the carbon dioxide is the *oxidizing agent* and becomes reduced.

The Can Ripper

An aluminum soda can is weakened as the result of a redox reaction. Select a 'sturdy' looking student from the class. Give the student an untreated pop can; ask the student to 'tear the can in half.' Then, select a student of small stature whom one would not suspect to be capable of such a feat. Give the student a treated can, and request that the same act be performed. A properly treated can is as easy to tear as is a piece of aluminum foil used for cooking.

Materials

Aluminum soda can
Copper(II) chloride, CuCl_2 , 2 g
Tap water

Hazards

Unsuspecting persons may accidentally drink from the can during the treatment. Do not allow this. The edges of the torn can may cause cuts.

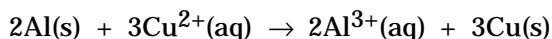
Precautions

Keep the can stored in a place where students cannot accidentally use it during the treatment. Wear gloves when 'ripping' the can.

Procedure

Score the *inside* of an aluminum soda can by inserting a triangular file into the can opening and scratching a line inside around the can (about 5 cm from the bottom). Place 1-2 g copper chloride in the can, and add enough water to cover the scoring. Swirl to dissolve the copper(II) chloride. Wait about 3-5 min. (Experiment with the time to avoid "eat through" of the can.) Pour the solution into a heavy metal waste container. Rinse the can. Give a student garden gloves, and ask the student to tear the can. (Chloride ion is needed for this procedure to work effectively. Copper(II) sulfate is not a suitable substitute for copper(II) chloride unless a chloride source such as NaCl or HCl also is added.)

Reaction



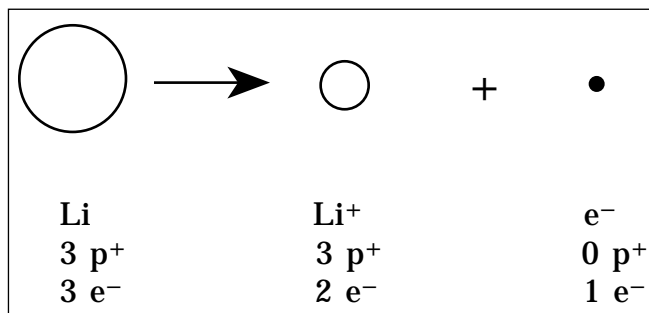
The copper(II) ions oxidize the aluminum metal.

Metaphors, Analogies, and Mnemonics

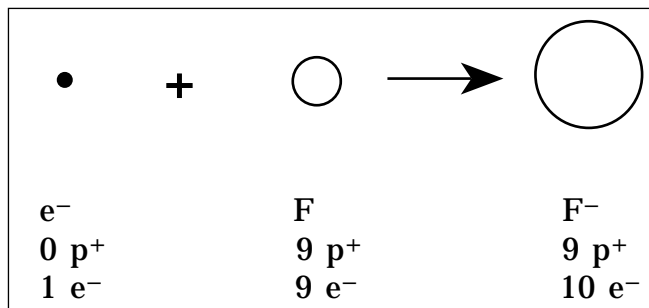
- Two excellent memory devices are available to help students keep redox terms in mind. Give a hearty growl as you walk into the classroom to introduce "LEO the lion says GER" (Loss of Electrons is Oxidation...Gain of Electrons is Reduction). Or have one of your art students make a colorful poster to place on the chemistry room wall showing LEO in full majesty saying GER.
- A second equally effective memory device is "OIL RIG" (Oxidation Is Loss...Reduction Is Gain). Again, a poster of an oil well spouting with the above statement included makes a great classroom decoration.

Pictures in the Mind

- Draw a picture of an atom being oxidized.

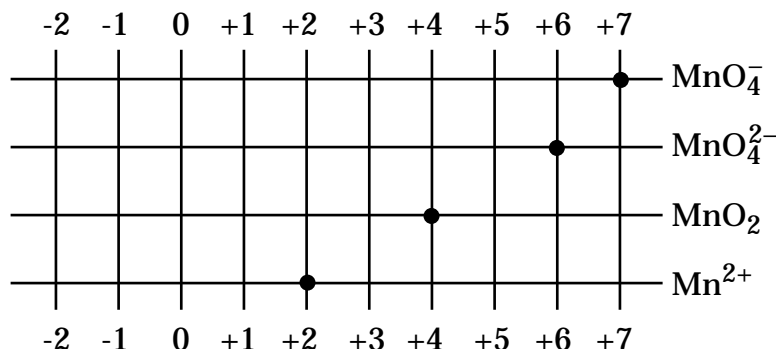


- Draw a picture of an atom being reduced.

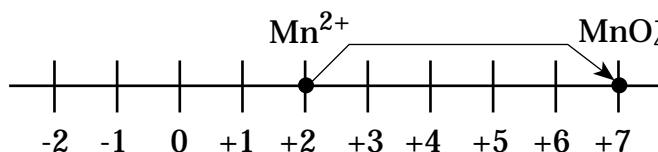




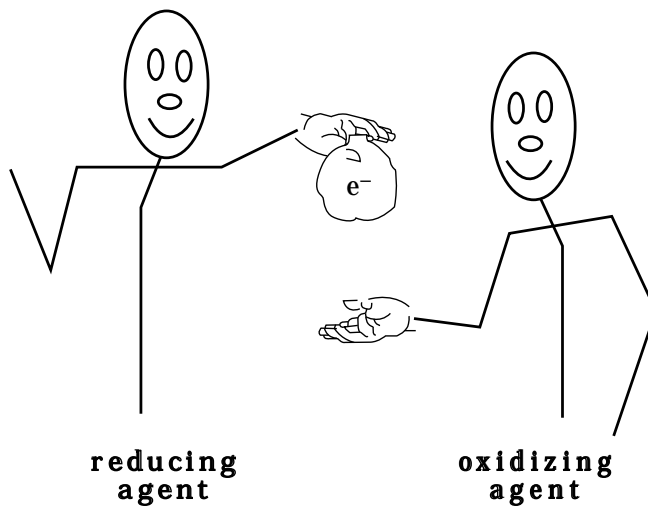
3. Using number lines, illustrate what happens to an atom that is reduced.



4. Illustrate movement on a number line to indicate oxidation.



5. Show a picture of an oxidizing agent at work.
6. Show a picture of a reducing agent at work.



coffee pot
filling agent
is emptied



coffee cup
emptying agent
is filled

TIPS FOR THE TEACHER

Language of Chemistry

combustion chemical reaction accompanied by heat and light, usually associated with oxidations using oxygen as the oxidizing agent.

corrosion reaction of metals forming compounds, particularly when exposed to air and water (to corrode is to eat away at or destroy slowly); metallic corrosion involves oxidation of the metal.

electron negatively charged subatomic particle with low mass.

oxidation reaction in which oxidation number increases (electrons are lost to another species).

oxidation number number used in keeping track of the change in distribution of electrons during reaction; it is assigned on the basis of several rules such as the oxidation number of an atom in its elemental form is zero.

oxidation state (see oxidation number)

oxidize to increase in oxidation number; to lose electrons to another species.

oxidizing agent chemical entity that brings about oxidation and becomes reduced; an entity that gains or picks up electrons during chemical reaction.

redox acronym used to describe reactions in which oxidation numbers change.

reduce to decrease in oxidation number; to gain electrons from another species.

reducing agent chemical entity that brings about reduction and becomes oxidized; an entity that loses electrons during chemical reaction.

reduction reaction in which oxidation number decreases (electrons are acquired from another species).

respiration biochemical reactions that provide energy through the metabolism or breakdown of biochemicals.

Oxidizing and Reducing Agents

After your students are comfortable with the concepts of oxidation and reduction (with the help of *LEO says GER* or *OIL RIG*), the concept of oxidizing agents and reducing agents should be introduced. Ask the class to define the term agent. Connect your discussion with previous experiences in earth science where students may have learned about “agents of erosion.” Ask them to list the agents of erosion. When they tell you wind, water, *etc.*, ask them why these are the “agents” of erosion. They will most certainly state that they are agents because they “cause” erosion.

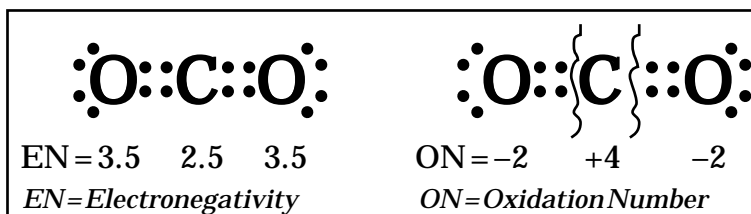
Tie this concept of an agent as a “causer” to the concept of oxidizing/reducing agents. An oxidizing agent is a “causer” of oxidation. Since oxidation is loss of electrons (*OIL*), an oxidizing agent *causes loss of electrons*. It should follow that if the oxidizing agent causes the loss of electrons by another substance, it is *gaining electrons*. The next logical step in the discussion is to help students see that if the oxidizing agent is gaining electrons, it is being reduced (*RIG*).

Correspondingly, a reducing agent must be a “causer” of reduction. Since reduction is gain of electrons (*RIG*), a reducing agent causes another substance to gain electrons. The only way it can cause gain of electrons is to lose electrons. If the reducing agent is losing electrons, it is being oxidized, since oxidation is loss (*OIL*).

To bring the discussion to closure, ask students to explain why a substance cannot be an oxidizing agent and, in the same reaction, be oxidized. If they understand the concepts involved, they will quickly see why oxidizing agents are not oxidized and why reducing agents are not reduced.

Assigning Oxidation Numbers

Oxidation number can be thought of as the charge a bonded atom would have if all bonding electrons were assigned arbitrarily to the more electronegative atom.





Oxidation Numbers are Assigned Using Several Arbitrary Rules

1. The oxidation number of an atom in its elemental state is 0. [S_8 , P_4 , O_2]
2. The oxidation number of an atom in a monatomic ion is equal to the charge on the ion. [$Fe^{3+} = +3$, $O^{2-} = -2$, $N^{3-} = -3$]
3. The oxidation number of fluorine atoms in a compound is -1 . [CF_4 , $F = -1$]
4. The oxidation number of alkali metal atoms in compounds is $+1$. [$Na^+ = +1$; $K^+ = +1$]
5. The oxidation number of alkaline earth metal atoms in compounds is $+2$. [$Ca^{2+} = +2$; $Mg^{2+} = +2$]
6. The oxidation number of aluminum in compounds is $+3$.
7. When halogen elements are in compounds with less electronegative elements, halogen atoms have oxidation numbers of -1 . [Cl in $NaCl = -1$; Cl in $PCl_3 = -1$; Cl in $HCl = -1$]
8. Hydrogen has an oxidation number of $+1$ with more electronegative elements. [H in $HI = +1$]
9. Oxygen usually has an oxidation number of -2 in compounds.
10. The sum of oxidation numbers of all atoms in a neutral molecule is 0. The sum of oxidation numbers of all atoms in an ion is equal to the charge of the ion.

Other Rules *(Not always taught in high school chemistry)*

1. The oxidation number of an atom in a (homonuclear) polyatomic ion is equal to the charge on the ion divided by the number of atoms in the ion. [$Hg_2^{2+} = +1$, $I_3^- = -1/3$]
2. Hydrogen has an oxidation number of -1 with less electronegative elements. [H in $NaH = -1$]
3. When combined with fluorine, oxygen has a positive oxidation number [O in $OF_2 = +2$]; in peroxides, oxygen has an oxidation number of -1 [O in $H_2O_2 = -1$; O in $BaO_2 = -1$]; oxygen has an oxidation number of $-1/2$ in superoxides [O in $KO_2 = -1/2$].

Using the Algebraic Number Line with Redox Equations

Students may have difficulty determining the number of electrons lost or gained in a reaction. They generally have difficulty deciding which substance involved is oxidized and which is reduced. The number line concept they may have learned in elementary algebra can be utilized to good advantage here.

A move to the right on a number line leads to a more positive number; a move to the left leads to a more negative number. In redox chemistry, since electrons with negative charges are being “moved around,” the concept is best illustrated by stating both moves in terms of negatives; that is, more positive can be viewed as being less negative. [An interesting aside is that this same process operates in electronics where, for example, a potential of $+5$ V can be considered to be negative with respect to a potential of $+10$ V. Since it is 5 V less positive, it is, in effect, 5 V more negative. When both points where voltmeter leads are to be connected are above ground potential (positive), electronic experimenters soon learn to connect the black (negative) lead from the voltmeter to the less positive (more negative) point.]

For a redox example, imagine that Fe^{3+} ions are converted to Fe^{2+} ions. Have students draw a simple number line to illustrate this oxidation number change. The number line move is from +3 to +2 (to the left). The oxidation number is now less positive (more negative) by one charge. The ion must have gained one electron to become one charge more negative. The number is *reduced* in magnitude by a move to the left. This fact will assist students in associating a move to the left on the algebraic number line with chemical reduction.

If I^{7+} is converted to I^- , there is a move of eight units to the left on the number line. The number is obviously less positive (more negative). Again, it is apparent that 8 electrons have been “gained” by the ion. Iodine has been “reduced” in both algebraic and chemical senses.

If one reverses the process so I^- is converted to I^{7+} , then the move is eight to the right on the number line. The resulting ion is eight charges more positive (less negative) than I^- . The only way an ion can become less negative is to lose electrons. The original ion must have been oxidized (*OIL*) in the chemical reaction. Encourage your students to use the algebraic number line concept to aid their understanding of redox concepts.

Pattern Recognition

1. More active (lower electronegativity; lower left-hand portion of Periodic Table) metals replace (reduce) less active (higher electronegativity) metals from a compound.
2. More active (higher electronegativity) nonmetals replace (oxidize) less active (lower electronegativity) nonmetals from a compound.
3. Metals reduce nonmetals. Nonmetals oxidize metals.
4. As an atom's oxidation state moves left on the number line, the species is being reduced. As its oxidation state moves to the right on the number line, the species is being oxidized.

Common Student Misconceptions

1. **“All the bonds formed are ionic, as shown by the use of oxidation numbers.”**

Oxidation numbers are assigned to atoms in molecules or ions that are described as being made up from essentially covalent bonds.

2. **“Oxidation number represents the actual charge of the atom in a molecule.”**

This is not so. For example, the iodine atom in IO_4^- is not thought of as a 7+ cation surrounded by four 2- anions, but instead as having electron pairs shared between the iodine and each oxygen atom.

3. **“All reactions are redox reactions.”**

Acid base reactions are not redox reactions; precipitation reactions such as $\text{BaSO}_4(\text{s})$ from $\text{Ba}^{2+}(\text{aq})$ and $\text{SO}_4^{2-}(\text{aq})$ are not redox reactions.

4. **“Oxidation always involves oxygen.”**

Oxidation involves a loss of electrons. Oxygen atoms are good at gaining electrons, but so are fluorine atoms, chlorine atoms, *etc.*



5. **“Electrons are lost when something is ‘reduced.’”**

Electrons are ‘lost’ when captured by certain radioactive nuclei or when they react in negative electron/positive electron pairs. Electrons are not lost (*i.e.*, are not consumed) in redox reactions. Instead, they may be thought of as moving from one species (the reducing agent) to another (the oxidizing agent).

6. **“A half reaction can occur by itself, *e.g.*, oxidation without reduction.”**

In real systems, reduction and oxidation are paired. The electrons required to bring about a reduction come from an oxidation. If an oxidation were to occur by itself without accompanying reduction, the electrons would still need to go someplace—and extra electrons, in large numbers, repel one another with enormous force.

7. **“All oxidation numbers must be integers.”**

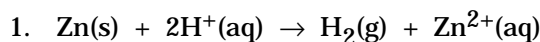
Oxidation number rules sometimes lead to fractional oxidation numbers. For example, the oxidation number of oxygen in potassium superoxide (KO_2) is $-1/2$ (see *Other Rules in Language of Chemistry*).

Problem Solving

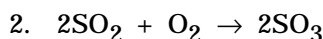
Students should:

1. Identify redox reactions from given equations.
2. Identify what is oxidized and what is reduced in a redox reaction.
3. Identify oxidizing and reducing agents in a redox equation.
4. Use rules to assign oxidation numbers to atoms in molecules, ions, and salts.
5. Balance simple redox equations when given the redox rules while keeping track of electron loss and gain.

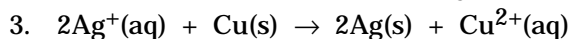
Samples



Zn, the reducing agent, is oxidized. The oxidation number of Zn changes from 0 to +2. H^+ , the oxidizing agent, is reduced. Its oxidation number changes from +1 to 0. The oxidation numbers are assigned as 0 for elements and are set equal to the charges for monatomic ions.



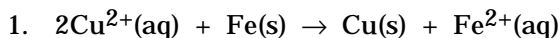
SO_2 , the reducing agent, is oxidized. The oxidation number of S changes from +4 to +6. O_2 , the oxidizing agent, is reduced. Its oxidation number changes from 0 to -2 . The oxidation number of O in O_2 is 0 (it is an element). The oxidation number of S in SO_2 is +4. Oxygen is assigned -2 in compounds. Two times -2 is -4 . Since the molecules have no net charge, this -4 must be balanced by the sulfur, so sulfur’s oxidation number must be +4. Similarly, the oxidation number of S in SO_3 is +6.



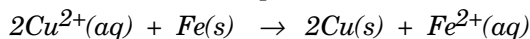
$\text{Ag}^+(\text{aq})$, the oxidizing agent, is reduced. Cu(s) , the reducing agent, is oxidized. One mole of Cu(s) is oxidized by two moles of $\text{Ag}^+(\text{aq})$.

Additional Examples

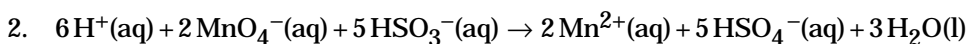
For each reaction depicted, identify the oxidizing agent, the reducing agent, the species oxidized, and the species reduced. Indicate whether the equation is balanced. Give the oxidation numbers of all atoms other than hydrogen and oxygen.



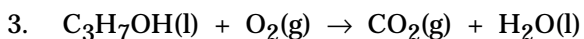
[Equation unbalanced; balanced equation is:



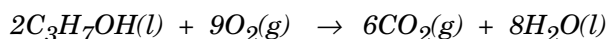
$\text{Cu}^{2+}(\text{aq})$ is oxidizing agent and is reduced; $\text{Fe}(\text{s})$ is reducing agent and is oxidized. Oxidation numbers: Cu in $\text{Cu}^{2+}(\text{aq})$, +2; Fe in $\text{Fe}^{2+}(\text{aq})$, +2; Cu in $\text{Cu}(\text{s})$, 0; Fe in $\text{Fe}(\text{s})$, 0.]



[Equation balanced. MnO_4^{-} is oxidizing agent and is reduced; HSO_3^{-} is reducing agent and is oxidized. Oxidation numbers: Mn in MnO_4^{-} , +7; S in HSO_3^{-} , +4; S in HSO_4^{-} , +6; Mn in Mn^{2+} , +2.]



[Equation unbalanced. Balanced equation is:



O_2 is oxidizing agent and is reduced; $\text{C}_3\text{H}_7\text{OH}$ is reducing agent and is oxidized. Oxidation numbers: C in $\text{C}_3\text{H}_7\text{OH}$, -2; C in CO_2 , +4.]

1. **Luigi Galvani (1737 - 1798).** During the 1770s Galvani began research in the investigation of animal electricity in nerves and muscles. From 1780 to 1791, he did extensive research into the effects elicited by static electricity on properly prepared frogs. The preparation consisted of removing the spinal cord, crucial nerves and lower limbs of the frog. Galvani then conducted experiments examining the convulsive contractions of the muscles in the lower limbs. He soon discovered an unexpected result: the muscles contracted without electrical stimuli when the frog, placed on an iron plate, was prodded with brass hooks. Galvani had discovered the central phenomenon of galvanism and produced an example of a redox reaction; the production of electric current was possible through the contact of two different metals in a moist environment. However, Galvani did not recognize the redox reaction; instead he interpreted his results as evidence for a theory that animal muscles possess their own unique electrical fluid.
2. **Alessandro Guiseppe Antonio Anastasia Volta (1745 - 1827).** The first accurate analysis of the phenomenon observed by Galvani was provided by Volta in 1792. His research, using the whole frog, showed that the muscle discharge was the result of electricity produced by dissimilar metals coming in contact with one another through a moist environment. Volta demonstrated first that galvanic excitations arose from external electrical stimulation and, secondly, that metals could be ranked by their electromotive power. In 1793 he demonstrated that the electromotive power of a chain of dissimilar metals depended only on the nature of the two extreme links, those touching the moist conductor. Volta went on to measure the voltage of a variety of metals undergoing redox reactions. Volta debated these theories with Galvani during the early 1790s. Volta's explanations for the electromotive power of metals and their production of voltage eventually proved to be a more accurate description of the redox phenomenon than that of the galvanists.
3. **Humphry Davy (1778 - 1829).** While Volta recognized that electricity was generated when two different metals were placed in solution, it was Humphry Davy who explained the chemical nature of this reaction. After reading Volta's work, Davy began working on electrolytic cells and demonstrated that a chemical reaction was generating the electricity.

HISTORY: ON THE HUMAN SIDE



4. **Hermann Walter Nernst (1864 - 1941).** Research on thermoelectricity provided the basis for Nernst's development of an electron theory of thermal and electrical conductivity. By examining solutions of various ionic concentrations separated by semipermeable partitions, Nernst soon demonstrated the relationship between ionic mobility, the diffusion coefficient and the electromotive force in concentration cells. As a result, he developed a mathematical equation for the electromotive force (E) in a galvanic process and a concentration change (C_1 to C_2):

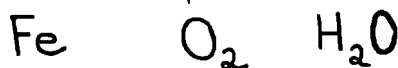
$$E = \frac{RT}{nF} \ln \frac{C_1}{C_2}$$

In this equation, R is the gas constant, T the absolute temperature, N the gram equivalents that have reacted, and F is Faraday's constant. Through this equation, a fundamental link between thermodynamics and electrochemical solution theory was established.

HUMOR: ON THE FUN SIDE

1. Student responses on exam:
 - a. Reduction is burning wood giving off different elements and substances.
 - b. Combustion is where there are a lot of greasy rags piled together and the air can't get through them. The air can go just so far and then it starts back. The air going in and coming out heats the rags and the rags get hot. Then they burst into flame.
- 2.

Wanna rust?



CHEM 13 NEWS, January 1983, p. 3

3. Word Search (see Appendix for master copy)

R O E Q C E F N O I T A D I X O A
L D A G E N T R K Y L Z F H W L N
R B N Y U B Q X X N V N N X K L J
C F O T L D X M Q H B F F A S V W
N G R Z D S Y I I D H X L M G I S
F W T S T D T R W R E I V Y I V V
L A C F S C A C O G M R D W R K E
U D E N U W R F K E G O Y P L X L
O X L D O R U M T E M K C K I A Q
R M E H R H C A G T A Z E R O W F
I R T M J U L J W G F K D J A O C
N R E B M U N N O I T A D I X O Y
E Z W P J A T D O I U P K P G H L

Words about the concepts in this module can be obtained from the clues given. Find these words in the block of letters:

1. Loss of electrons.
2. Gain of electrons.
3. The entity that causes an event to occur.

4. Mnemonic used to remember which species gains and which loses electrons.
5. Charge a bonded atom would have if all bonding electrons were assigned to the more electronegative atom. (2 words)
6. Most highly electronegative element.
7. The sum of oxidation numbers of all atoms in a neutral molecule.
8. This family of atoms in compounds has an oxidation number of +1. (2 words)
9. Shortened form of writing oxidation-reduction.
10. Entity transferred in an oxidation-reduction reaction.

Answers: 1. OXIDATION 2. REDUCTION 3. AGENT 4. OILRIG
 5. OXIDATION NUMBER 6. FLUORINE 7. ZERO 8. ALKALI METAL
 9. REDOX 10. ELECTRON

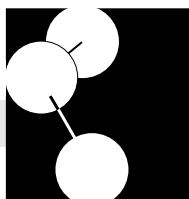
4. See cartoons at end of module.

MEDIA

1. *The World of Chemistry* videotape "Number 15: The Busy Electron," can be used by chemistry teachers at all levels. World of Chemistry Videocassettes. Annenberg/CPB Project, P.O. Box 1922, Santa Barbara, CA 93116-1922; (800) 532-7637; World of Chemistry Series, Atlantic Video, 150 South Gordon Street, Alexandria, VA 22304; (703) 823-2800 or QUEUE Educational Video, 338 Commerce Drive, Fairfield, CT 06430; (800) 232-2224.
2. *The World of Chemistry* (high school version) videotape "Number 17: Oxidation and Reduction," WINGS for Learning/SUNBURST, 101 Castleton Street, Pleasantville, NY 10570; (800) 321-7511; (914) 747-3310; (914) 747-4109 (FAX).
3. CHEM Study films/videos available from Ward's Natural Science Establishment, Inc., P.O. Box 92912, Rochester, NY 14692-9012; (800) 962-2660.
 - a. Volume 24: *Vanadium, A Transition Element*
 - b. Volume 20: *Nitric Acid*
4. *Doing Chemistry*, videodisc set available from the American Chemical Society, 1155-16th St., N.W., Washington, DC 20036; (202) 872-4382.
5. *Closeup on Chemistry*, video available from the American Chemical Society, 1155-16th St., N.W., Washington, DC 20036; (202) 872-4382.
6. Software published by *JCE: Software*, a publication of the *Journal of Chemical Education*, Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue, Madison, WI 53706-1396: (608) 262-5153 (voice) or (608) 262-0381 (FAX).
 - a. *REDOX: A Tutorial on REDuction/OXidation Equations*, by Derek Davenport, Paul Groves and Dale Jensen. Vol. III A, No. 1 for the Apple II computer.
 - b. *Oxidation Number Rules*, by James P. Birk. Vol. VI B, No. 1, for IBM PS/2, PC-compatible computers.
 - c. *Frost Diagrams*, by James P. Birk and Heidi Hocker. Vol. IV C, No. 1, for the Apple Macintosh.



7. Videodiscs published by *JCE: Software*, a publication of the *Journal of Chemical Education*, Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue, Madison, WI 53706-1396; (608) 262-5153 (voice) or (608) 262-0381 (FAX).
 - a. "Copper-Zinc Electrochemical Cell," a chapter on *The World of Chemistry: Selected Demonstrations and Animations: Disc I* (double sided, 60 min.), Special Issue 3.
 - b. "The Aufbau Principle," "Making Sodium Chloride" and "Thermite Reaction," three chapters on *The World of Chemistry: Selected Demonstrations and Animations: Disc II* (double sided, 60 min.), Special Issue 4.
 - c. *The Periodic Table Videodisc* (single side, 30 min.), Special Issue 1.
 - d. "Combustion of Methane," "Oxidation of Alkanes, Alkenes and Alcohols," and "Acetylene – Preparation and Reactions," on the videodisc *Demonstrations in Organic Chemistry* (double sided, 60 min.), Special Issue 6.
8. *Redox*, by Helen Brooks and David Brooks. Videodisc distributed by Project SERAPHIM, Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue, Madison, WI 53706-1396; (608) 263-2837; (608) 262-0381 (FAX).



Links/Connections

1. **Electrochemistry.** Electrochemistry involves oxidation or reduction at the interface between a metallic conductor or semiconductor (an electrode) and an electrolyte. All electrochemical reactions are redox reactions.
2. **Stoichiometry.** Redox reactions are associated with changes of whole numbers of electrons (there are no fractions of electrons), so redox reactions involve substances in ratios of whole numbers as do other stoichiometric reactions.
3. **Periodicity.** The Periodic Table permits prediction of redox reactivity patterns of elements.
4. **Thermodynamics.** Most nonredox reactions studied (acid/base; lattice formation) have the energy changes accompanying the reactions limited in some special way by the nature of the reaction. (For example, aqueous solutions of acids and bases react to produce water. That main reaction, largely independent of particular acid and/or base sources, involves hydrogen ion and hydroxyl ion. The reaction $[H^+ + OH^- \rightarrow H_2O]$ limits the quantity of energy released. This energy limit is much smaller than is the limit for energetic redox reactions such as the formation of water from hydrogen and oxygen.)

WITHIN CHEMISTRY

1. Biochemistry (majority of biochemical reactions are redox—*e.g.*, cell respiration.)
2. Food Science and Food Preservation
3. Geochemistry—*e.g.*, nitrogen cycle
4. Atmospheric Sciences—*e.g.*, photochemical smog (Nitrogen gas, N_2 , is one of the most stable substances known, but when subjected to high temperatures as in internal combustion engines, it can be oxidized by atmospheric oxygen, O_2 , to nitrogen(II) oxide, NO , and nitrogen(IV) oxide, NO_2 . Nitrogen(IV) oxide is responsible for much of the “brown cloud” visible as smog in large urban areas.)
5. Civil Engineering—*e.g.*, corrosion prevention on infrastructures

BETWEEN CHEMISTRY AND OTHER DISCIPLINES

Personal

1. Corrosion of the Statue of Liberty (see *ChemMatters*, April, 1985)
2. The sulfur from an egg smashed on an automobile may bring about car paint damage through redox reactions
3. Antioxidants in food
4. Decaying food
5. Polluted water caused by fossil fuel and high-sulfur coal burning. (Chemical oxidation of reduced carbon compounds—burning of coal, petroleum, wood—remains the major source of energy needed to maintain our lifestyles. Because much organic material contains sulfur as well as carbon, many air pollutants [SO_2 , SO_3 , CO , CO_2 , NO , and NO_2] are the direct result of this energy production.)

TO THE CONTEMPORARY WORLD

Community

Field Trips

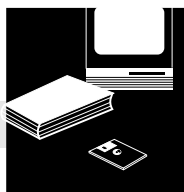
1. Local coal or gas fueled power plant
2. Electroplating factory
3. A metal producing factory (smelter, blast furnace, *etc.*)

Experts

1. Mining engineers
2. Chemists at electric power plants

Societal

1. **Decaying of metallic structures.** Corrosion, a major pathway of decay, requires replacement of existing structures made of steel.
2. **Photochemical smog and ozone layer.** Redox reactions lead to pollutants in the atmosphere. Some of these cause smog formation at ground level, while others destroy the ozone layer above.
3. **Greenhouse effect.** Although the oxygen quantity in the atmosphere is too extensive to be critically affected by human activities, the concentration of minor atmospheric gases can be changed drastically. Current concern about the greenhouse effect due to extensive oxidation of fossil fuels is a result of human activity. Even though it is a common gas we often mention when asked to name components of the atmosphere, CO_2 is present in only small amounts. The extent of fossil fuel burning is so great that the amount of CO_2 in the atmosphere is undergoing a significant percent increase. This gas prevents infrared radiation from leaving the earth and has the net effect of increasing the earth's temperature.
4. **Acid rain.** Sulfur in coal burns (oxidizes) to form sulfur dioxide. Sulfur dioxide is further oxidized to sulfuric acid in a series of steps including both oxygen and water. As the result of burning large amounts of coal, acid rain has played havoc with the ecological balance of many areas of the United States and Canada, especially in areas containing little or no limestone.
5. **Aging.** One of the principal causes of aging in organisms is attributed to effects of oxidation by 'radicals' formed when oxygen reacts in a biological system. Organisms have complex systems designed to prevent and to overcome the effects oxidative aging. On a different level, the freshness of nonliving food materials is often extended by adding substances called antioxidants. A simple example is the addition of vitamin C (ascorbic acid) when canning fruits to prevent browning (caused by air oxidation of the amino acid tyrosine and other phenolic materials followed by polymerization of the oxidation products).



References

Module developed by David Brooks, Robert Curtright, and James McGahan, the Nebraska team.

Banks, A. *Periodic table*. [Videodisc and accompanying written materials]. Madison, WI: Journal of Chemical Education Software. (Distributed by Project SERAPHIM, Department of Chemistry, University of Wisconsin, Madison, WI 53706)

Brooks, D. W. (Producer). (1989). *Doing chemistry*. [Videodiscs, computer program, and supporting written materials]. Washington, DC: American Chemical Society

Brooks, H. B., and Brooks, D. W. (Producers). *Redox*. [Videodisc, computer program, and written materials]. Washington, DC: American Chemical Society. (Distributed by Synaps, 334 South Cotner Blvd., Lincoln, NE 68510.)

Kemp, M. (1981). Tested demonstrations: Silver mirror. *Journal of Chemical Education*, 58, 655-656.

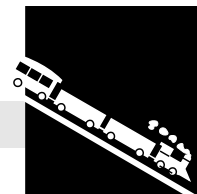
The Woodrow Wilson microscale version of Laboratory Activity 2: The Silver Mirror Reaction was developed by Fen Lewis, Strongville High School, 7701 Beach Road, Wadsworth, OH 44281, and was based upon this article.

Morgan, M. (1993, January). Ideas from the coast: Redox. *Connecting with Chemistry (Project SERAPHIM News)*, p. 3, 9.

Shakhashiri, B. Z. (1989). *Chemical demonstrations: A handbook for teachers of chemistry* (Vol. 3). Madison, WI: University of Wisconsin Press.

Summerlin, L. R., and Ealy, J. L. (1985). *Chemical demonstrations: A sourcebook for teachers*. Washington, DC: American Chemical Society

Appendix



- **Transparency Masters**
 1. Oxidation and Reduction—Redox
 2. Meet the Agents
 3. Word Search
- **Humor**

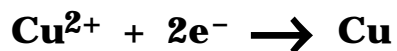
Oxidation and Reduction — Redox

Oxidation is a loss of electrons:



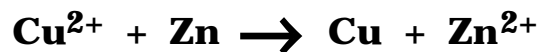
Zinc is oxidized.

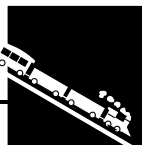
Reduction is a gain of electrons:



Copper is reduced.

Oxidation and reduction always occur together:





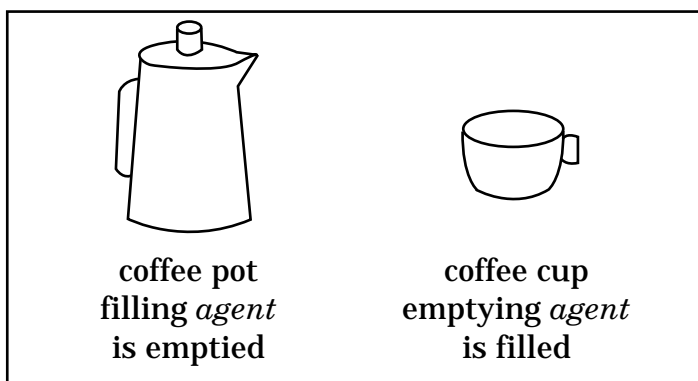
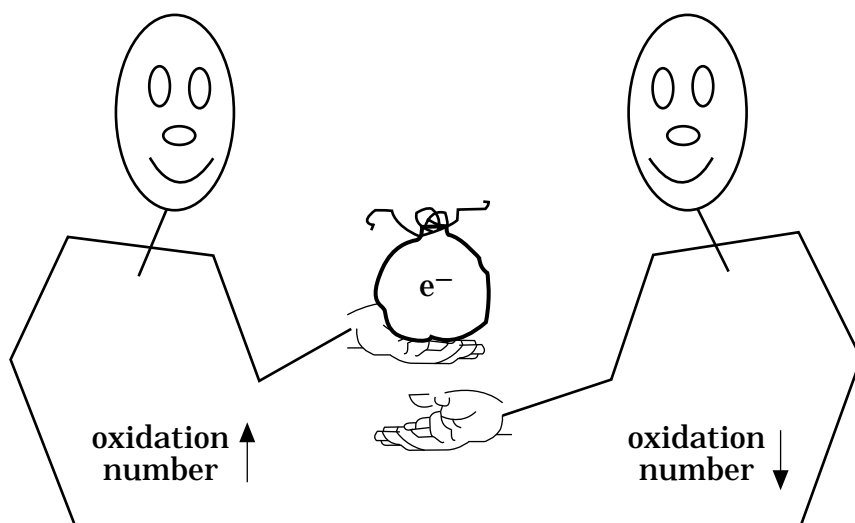
MEET THE Agents

REDUCING

*(loses electrons,
becomes oxidized)*

OXIDIZING

*(gains electrons,
becomes reduced)*

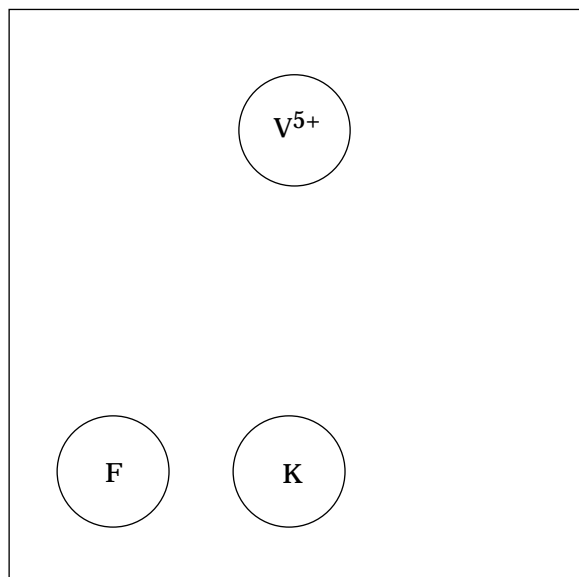
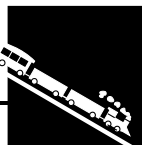


Word Search

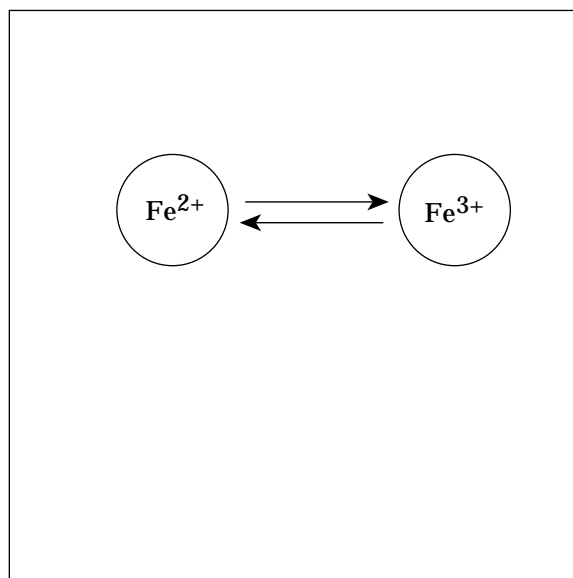
R O E Q C E F N O I T A D I X O A
L D A G E N T R K Y L Z F H W L N
R B N Y U B Q X X N V N N X K L J
C F O T L D X M O H B F F A S V W
N G R Z D S Y I I D H X L M G I S
F W T S T D T R W R E I V Y I V V
L A C F S C A C O G M R D W R K E
U D E N U W R F K E G O Y P L X L
O X L D O R U M T E M K C K I A Q
R M E H R H C A G T A Z E R O W F
I R T M J U L J W G F K D J A O C
N R E B M U N N O I T A D I X O Y
E Z W P J A T D O I U P K P G H L

Words about the concepts in this module can be obtained from the clues given.
Find these words in the block of letters:

1. Loss of electrons.
2. Gain of electrons.
3. The entity that causes an event to occur.
4. Mnemonic used to remember which species gains and which loses electrons.
5. Charge a bonded atom would have if all bonding electrons were assigned to the more electronegative atom. (2 words)
6. Most highly electronegative element.
7. The sum of oxidation numbers of all atoms in a neutral molecule.
8. This family of atoms in compounds has an oxidation number of +1. (2 words)
9. Shortened form of writing oxidation-reduction.
10. Entity transferred in an oxidation-reduction reaction.

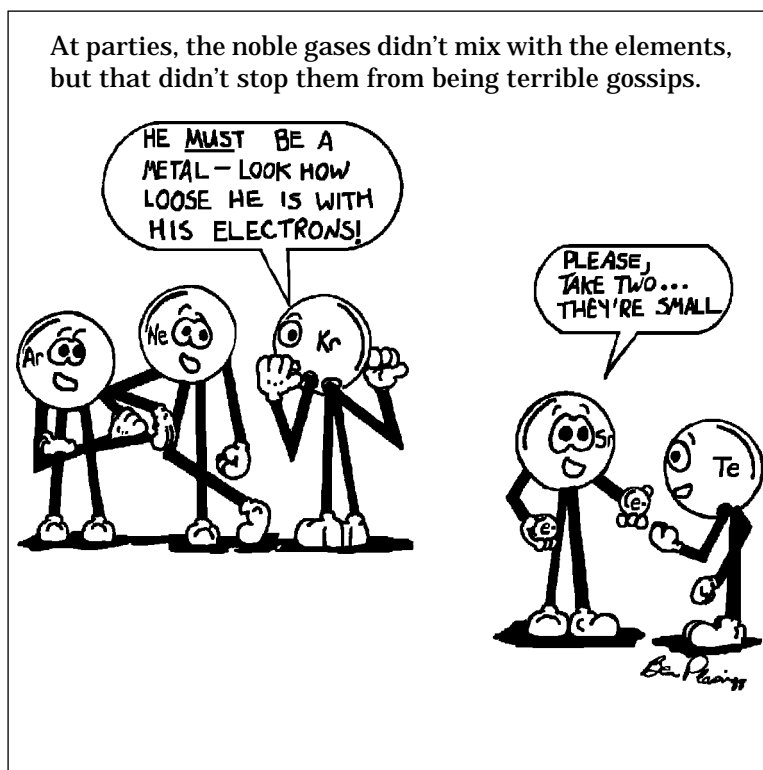


“Kay, the only problem with Van is you never know from day to day what state he’ll be in!”

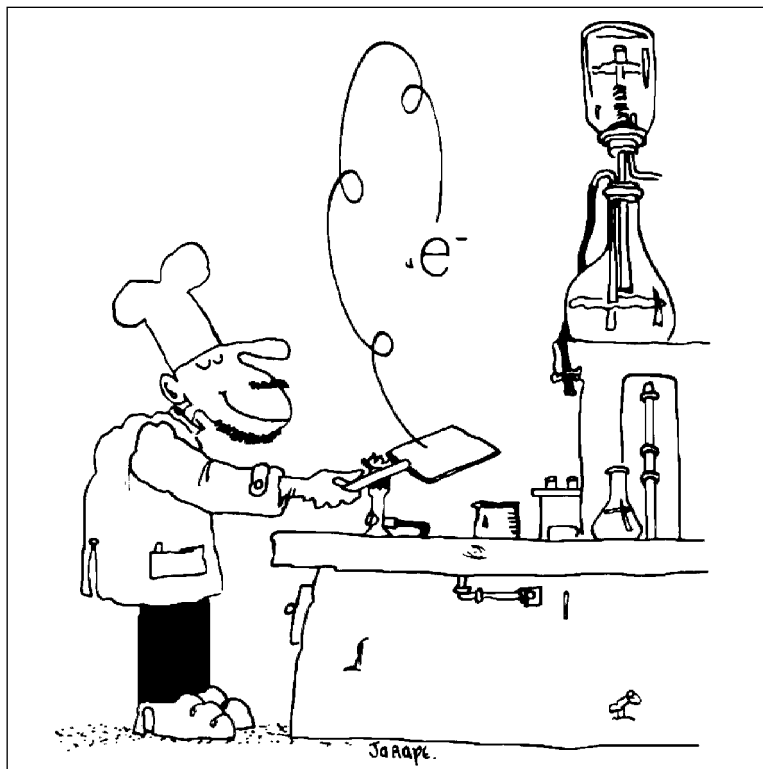
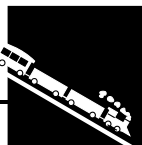


“Problem with Ferd is he can never seem to make up his mind.”

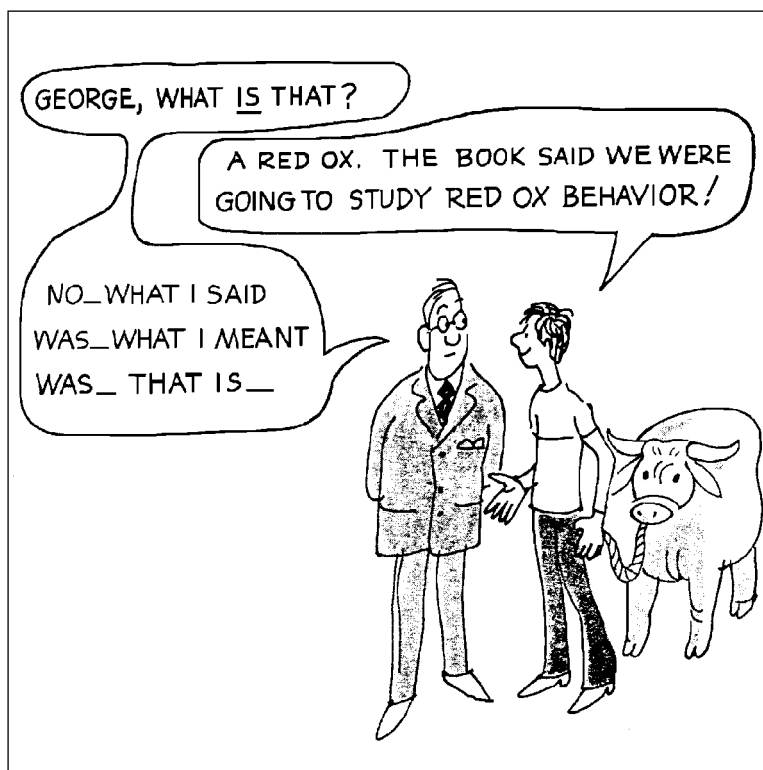
© John Ihde, Wausau West High School, Wausau, WI 54401.
Reprinted with permission.



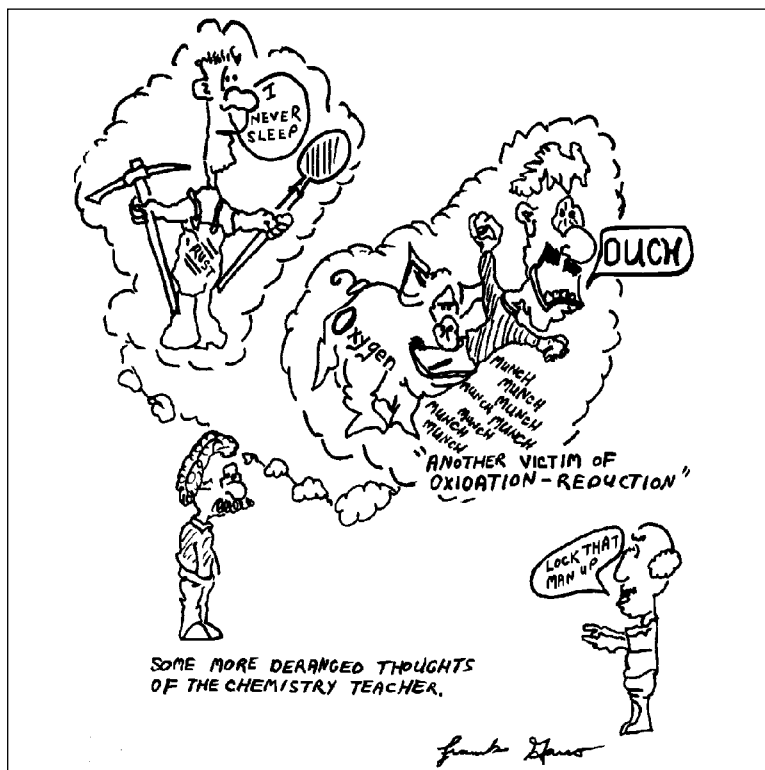
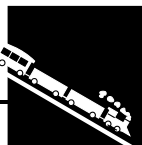
CHEM MATTERS, December 1988, p. 2. Reprinted with permission.



CHEM 13 NEWS, September 1988, p. 26. Reprinted with permission.



Used by permission of Veritex Publishing Co.



CHEM 13 NEWS, March 1981, p. 3. Reprinted with permission.