Process Oriented Guided Inquiry Learning in a High School Science Classroom: A Method for Teaching both Process Skills and Content Knowledge
Lesson Plan for EDUC 489 – Special Topics: Science Wise Course

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Title: An Introduction to Redox through POGIL

Subject Area(s): Chemistry

Grade(s): 11-12

Description of Lesson: The purpose of this lesson is to introduce students to chemical reactions that involved the simultaneous processes of oxidation and reduction as electrons are transferred from one chemical species to another. Students are placed in cooperative groups to process through the tasks outlined in the lesson. POGIL is predicated upon having small groups of students review one or more models prior to answering a series of questions related to the model(s). The instructor acts as facilitator to assist each group in filling in their knowledge base as they work through the POGIL process.

Length of Lesson: One 45-minute class period

Student Objectives: By the end of the lesson, students should be able to distinguish oxidation and reduction and to identify which chemical species act as oxidizing agents and which act as reducing agents.

Materials:
For the teacher: A copy of the POGIL activity to use in guiding student groups.
For each group: Copies of the POGIL activity and a periodic table for each student.
For the class: Copies of the POGIL activity and a periodic table for each student.

Procedure: Students are divided into groups of four and are assigned roles (spokesperson, timekeeper, person to check for understanding, person to maintain on-task group behavior, etc.) to use during the POGIL process. As groups analyze the model(s), they process through a series of guiding questions that scaffold their knowledge in order to develop both process and content skills.

Assessment: Upon completion of the Oxidation-Reduction POGIL, students will engage in an individual quiz (labeled Quiz 4.3 at the end of the lesson plan) so that their progress can be assessed more formally by the instructor in the event that additional review or follow-up activities are necessary.
Missouri and Kansas Standards Addressed:

**Kansas Science Standards:**
Standard 2A: Chemistry, Grades 8-12
The students will develop an understanding of the structure of atoms, *compounds*, chemical reactions, and the interactions of energy and matter.

Benchmark 3: The student will gain a basic concept of chemical reactions.
   The student ...
   1. understands a chemical reaction occurs when one or more substances (reactants) react to form a different chemical substance(s) (products). There are different types of chemical reactions all of which demonstrate the Law of Conservation of Matter and Energy.
      a. Chemical reactions are written as balanced chemical equations. In ordinary chemical reactions, the number and kind of atoms must be conserved.
      b. Examples of chemical reactions are synthesis, decomposition, combustion, single and double replacement, acid/base, and oxidation/reduction.

**Missouri Science Standards (GLE’s):**
Strand 1: Properties and Principles of Matter and Energy
1. Changes in properties and states of matter provide evidence of the atomic theory of matter
   1. Mass is conserved during any physical or chemical change
      Scope and Sequence – Atomic Theory and Changes in Matter
      a. Compare the mass of the reactants to the mass of the products in a chemical reaction or physical change as support for the Law of Conservation of Mass
      b. Recognize whether the number of atoms of the reactants and products in a chemical equation are balanced
The Chemical Reaction of Fe(s) and Cu^{2+}(aq)

When iron filings are placed in a 1.0 M copper(II) sulfate solution and left to stand for a while, copper is removed from solution and deposits on the bottom of the beaker as a red solid, and some Fe^{2+} ions are found in solution. Assuming iron is the excess reactant, eventually all of the copper ions will settle out as solid copper. This type of reaction, which involves a transfer of electrons is called an oxidation-reduction, or redox, reaction.

Equation 1:  
Fe(s) + Cu^{2+}(aq) → Fe^{2+}(aq) + Cu(s)

CRITICAL THINKING QUESTIONS:
1. Which reactant in Equation 1 . . . (circle one)
   a. Gains electrons?  Fe(s) or Cu^{2+}(aq)
   b. Loses electrons?  Fe(s) or Cu^{2+}(aq)

2. What is the charge on the solid iron atoms on the left side of the equation?
   What is the charge on the iron ions on the right side of the equation?

3. What is the charge on the aqueous copper ions on the left side of the equation?
   What is the charge on the solid copper atoms on the right side of the equation?

4. How many electrons are transferred when . . .
   a. One iron atom reacts with one Cu^{2+} ion?
   b. One mole of iron reacts with one mole of Cu^{2+}?

Oxidation and Reduction

Redox reactions are sometimes divided into half-reactions to separate and clarify the electron transfer process. The species that loses electrons is said to be oxidized and is often referred to as the reducing agent. The species that gains electrons is said to be reduced and is referred to as the oxidizing agent. Every redox reaction must have both a reducing agent (the species that loses electrons) and an oxidizing agent (the species that gains electrons.

Sometimes an acronym helps to remember these processes:
   LEO the lion goes GER  (LEO = Loss of Electrons is Oxidation and GER = Gain of Electrons is Reduction)
   OIL RIG  (Oxidation Is Loss, Reduction Is Gain)
Examine this redox reaction:

\[ 2 \text{Br}^1(\text{aq}) + \text{Cl}_2(\text{g}) \rightarrow \text{Br}_2(\text{aq}) + 2 \text{Cl}^1(\text{aq}) \]

The half reactions that describe the electron transfer process for this reaction are:

\[ 2 \text{Br}^1(\text{aq}) \rightarrow \text{Br}_2(\text{aq}) + 2 e^- \]
\[ \text{Cl}_2(\text{g}) + 2 e^- \rightarrow 2 \text{Cl}^1(\text{aq}) \]

CRITICAL THINKING QUESTIONS:

5. Which species becomes more positive during the reaction? \(2 \text{Br}^1(\text{aq})\) or \(\text{Cl}_2(\text{g})\)

Is this the reducing agent or the oxidizing agent? Explain.

Is this the species that is oxidized, or the one that is reduced? Explain.

6. Which becomes more negative? \(2 \text{Br}^1(\text{aq})\) or \(\text{Cl}_2(\text{g})\)

Is this the reducing agent or the oxidizing agent? Explain.

Is this the species that is oxidized, or the one that is reduced? Explain.

PROBLEMS:

7. Identify the reducing agent (RA) and the oxidizing agent (OA) in each of the following reactions:

a. \(\text{Br}_2(\text{aq}) + \text{Hg(s)} \rightarrow 2 \text{Br}^-(\text{aq}) + \text{Hg}^{2+}(\text{aq})\)

b. \(2 \text{Co}^{3+}(\text{aq}) + 2 \text{Br}^-(\text{aq}) \rightarrow \text{Br}_2(\text{aq}) + 2 \text{Co}^{2+}(\text{aq})\)

c. \(S_2\text{O}_8^{2-}(\text{aq}) + \text{Zn(s)} \rightarrow \text{Zn}^{2+}(\text{aq}) + 2 \text{SO}_4^{2-}(\text{aq})\)
d. \( 2 \text{Mg}(s) + \text{O}_2(g) \rightarrow 2 \text{MgO}(s) \)

e. \( 3 \text{Pb}^{2+}(aq) + 2 \text{Al}(s) \rightarrow 2 \text{Al}^{3+}(aq) + 3 \text{Pb}(s) \)

f. \( \text{Cu}^{2+}(aq) + \text{SO}_4^{2-}(aq) + \text{Zn}(s) \rightarrow \text{Zn}^{2+}(aq) + \text{Cu}(s) + \text{SO}_4^{2-}(aq) \)

**Oxidation Numbers**

Oxidation numbers are an accounting system for electrons. One of the main uses of oxidation numbers is to locate the oxidized and reduced species in a reaction if they are not immediately obvious. For example:

\[
\begin{align*}
2 \text{CuI}(s) & \rightarrow \text{Cu}(s) + \text{Cu}^{2+}(aq) + 2 \text{I}^{−}(aq) \\
\text{note: of these 2 Cu}^{+} ions, one is oxidized and one is reduced
\end{align*}
\]

The species that is oxidized (in this case, the Cu in CuI changes from +1 to +2) undergoes an increase in oxidation number – it becomes more positive.

The species that is reduced (in this case, the Cu in CuI changes from +1 to 0) undergoes a decrease in oxidation number – it becomes more negative.

Oxidation numbers are written above the atomic symbols in these examples:

\[
\begin{align*}
\text{HClO}_4 & : +1, +7 \text{each} \\
\text{O}_2 & : -2 \text{each} \\
\text{CH}_4 & : 0 \\
\text{Note that oxidation numbers for atoms, each individual atomic symbol's charge}
\end{align*}
\]

**Rules for Assigning Oxidation Numbers (ON)**

For atoms in elemental form, such as \( \text{Na}(s), \text{O}_2(g), \text{Cl}_2(g), \text{Br}_2(l) \) the ON = 0

For monatomic ions, such as \( \text{Mg}^{2+}, \text{O}^{2−} \) the ON = the ion charge

The sum of ON values for atoms in a molecule or formula unit = 0
The sum of ON values for atoms in a polyatomic ions = ion charge
ON for . . . Group 1A = +1
    Group 2A = +2
    Hydrogen = +1 if H is combined with a nonmetal, but -1 if H is combined with a metal
    Fluorine = -1
    Oxygen = -2 (except for the peroxide ion – in peroxide each O is -1)
Group 7A = -1 (except with O)

Notice that the oxidation numbers of bromine changes as the other atoms differ:
\[ \text{HBrO}_2 + S \rightarrow \text{HOBr} + \text{HOBrO} \]

Many transition elements and many member of Group 5A are likely to have multiple oxidation numbers, depending upon the compound in which they are found.

CRITICAL THINKING QUESTIONS:
8. Determine the oxidation number (ON) for the underlined atom in each of these compounds:

   a. \( \text{CrCl}_3 \)  
   b. \( \text{SO}_3 \)  
   c. \( \text{HNO}_3 \)  
   d. \( \text{SiO}_3 \)  
   e. \( \text{HPO}_4^{2-} \)  
   f. \( \text{IF}_3 \)  
   g. \( \text{H}_2\text{SO}_4 \)  
   h. \( \text{KMnO}_4 \)  
   i. \( \text{P}_4\text{O}_7 \)  
   j. \( \text{Br}_2\text{O} \)  
   k. \( \text{Br}_2 \)  
   l. \( \text{SO}_4^{2-} \)

9. Determine the oxidation number (ON) for each atom in these compounds:

   a. \( \text{H}_2\text{SO}_4 \)  
   b. \( \text{CH}_4 \)  
   c. \( \text{P}_4\text{O}_6 \)  
   d. \( \text{P}_4\text{O}_8 \)  
   e. \( \text{CrO}_4^{2-} \)  
   f. \( \text{Cr}_2\text{O}_7^{2-} \)  
   g. \( \text{NH}_2^{+} \)  
   h. \( \text{NO}_3^{-} \)

10. What is the oxidation number of the N and H in \( \text{NH}_3 \)? What is the oxidation number of Cu in \( \text{Cu}\left(\text{NH}_3\right)_4^{2+} \)?

11. What is the oxidation number of O and H in \( \text{OH}^{-} \)? What is the oxidation number of Al in \( \text{Al}\left(\text{OH}\right)_4^{1-} \)?

12. Which of these are redox reactions? (circle all that apply)

   a. \( 3 \text{H}_2(g) + \text{N}_2(g) \rightarrow 2 \text{NH}_3(g) \)
   b. \( \text{Cu}^{2+}(aq) + 4 \text{NH}_3(aq) \rightarrow \text{Cu}\left(\text{NH}_3\right)_4^{2+}(aq) \)
Balancing Redox Equations Using the Oxidation Number Method

It is essential to realize that the reducing agent loses electrons and the oxidizing agent gains them simultaneously. The transferred electrons are never free, which means we can balance a redox reaction by making sure that the number of electrons lost by the reducing agent equals the number of electrons gained by the oxidizing agent.

The oxidation number method is the method we will use to balance redox equations. It consists of five steps that use the change in oxidation numbers to generate the coefficients used to balance a redox equation. An example is shown below each step:

Step 1: Assign oxidation numbers to all elements in the reaction.

\[
\begin{align*}
\text{Cu} + \text{HNO}_3 & \rightarrow \text{Cu(NO}_3\text{)}_2 + \text{NO}_2 + \text{H}_2\text{O} \\
\text{Cu} & \quad \overset{+2}{\text{Cu(NO}_3\text{)}_2} + \overset{+1}{\text{NO}_2} + \overset{+2}{\text{H}_2\text{O}} \\
\text{Cu} & \quad \overset{0}{\text{Cu}} + \overset{+1}{\text{H}} \quad \overset{0}{\text{O}} \quad \overset{5}{\text{N}} \quad \overset{0}{\text{O}} + \overset{0}{\text{H}} \quad \overset{2}{\text{O}} \\
\text{Cu} & \quad \text{Cu} + \overset{5}{\text{N}} \quad \overset{0}{\text{O}} \quad \overset{0}{\text{O}} + \overset{0}{\text{H}} \quad \overset{2}{\text{O}} \\
\text{Note that the N in NO}_3^- & \text{ is always +5, no matter what to }
\end{align*}
\]

Step 2: Use the oxidation numbers to identify the oxidized and reduced species.

\[
\begin{align*}
\text{Cu} + \overset{+5}{\text{HNO}_3} & \rightarrow \overset{+2}{\text{Cu(NO}_3\text{)}_2} + \overset{+1}{\text{NO}_2} + \overset{+2}{\text{H}_2\text{O}} \\
\text{Cu} & \quad \overset{0}{\text{Cu}} + \overset{+1}{\text{H}} \quad \overset{0}{\text{O}} \quad \overset{5}{\text{N}} \quad \overset{0}{\text{O}} + \overset{0}{\text{H}} \quad \overset{2}{\text{O}} \\
\text{Cu} & \quad \overset{0}{\text{Cu}} + \overset{+1}{\text{H}} \quad \overset{0}{\text{O}} \quad \overset{5}{\text{N}} \quad \overset{0}{\text{O}} + \overset{0}{\text{H}} \quad \overset{2}{\text{O}} \\
\text{Step 3: Compute the number of electrons lost in the oxidation and those gained in the reduction from the oxidation number changes. Draw lines between these atoms to show the changes.}
\end{align*}
\]

\[
\begin{align*}
\text{Cu} + \overset{+5}{\text{HNO}_3} & \rightarrow \overset{+2}{\text{Cu(NO}_3\text{)}_2} + \overset{+1}{\text{NO}_2} + \overset{+2}{\text{H}_2\text{O}} \\
\text{Cu} & \quad \overset{0}{\text{Cu}} + \overset{+1}{\text{H}} \quad \overset{0}{\text{O}} \quad \overset{5}{\text{N}} \quad \overset{0}{\text{O}} + \overset{0}{\text{H}} \quad \overset{2}{\text{O}} \\
\text{Step 4: Multiply one or both of the numbers by appropriate factors to make the electrons lost equal the electrons gained, and use these as the balancing coefficients.}
\end{align*}
\]

\[
\begin{align*}
\text{Cu} + \overset{+5}{\text{HNO}_3} & \rightarrow \overset{+2}{\text{Cu(NO}_3\text{)}_2} + \overset{+1}{\text{NO}_2} + \overset{+2}{\text{H}_2\text{O}} \\
\text{Cu} & \quad \overset{0}{\text{Cu}} + \overset{+1}{\text{H}} \quad \overset{0}{\text{O}} \quad \overset{5}{\text{N}} \quad \overset{0}{\text{O}} + \overset{0}{\text{H}} \quad \overset{2}{\text{O}} \\
\text{Step 5: Complete the balancing by inspection.}
\end{align*}
\]

\[
\begin{align*}
2\text{HNO}_3 + 2\overset{+2}{\text{Cu}} & \rightarrow \overset{+2}{\text{Cu(NO}_3\text{)}_2} + \overset{+1}{\text{NO}_2} + \overset{+2}{\text{H}_2\text{O}} \\
\text{It would also be ok to change the 2HNO}_3 \text{ to 4... I added it separately to show the ox. and red. were separate from these.}
\end{align*}
\]
Example:

1. Each Cr gains 3 e− since there are 2 Cr on left; make 2 on right by adding coefficient. So 2 Cr’s lose 6 e− total.

2. Each I loses 1 e− since there are 2 on the right, make 2 on left; thus, 2 e− lost total.

3. e− lost must = e− gained. So × I by 3 so 6 e− total gained

\[ K_2Cr_2O_7(aq) + 6 \times \text{HI(aq)} \rightarrow 2 \times \text{Kl(aq)} + 2 \times \text{CrI}_3(aq) + 3 \times \text{I}_2(s) + 7 \times \text{H}_2\text{O(l)} \]

4. Balance K by making 2 KI

5. Balance O and H by making 7 H2O

13. Use the oxidation number method to balance the following reactions. They have been started for you:

Show all work (oxidation numbers, which species are gaining and which are losing electrons, etc.)

a. \[ \text{KClO}_3(aq) + \text{HBr(aq)} \rightarrow \text{Br}_2(l) + \text{H}_2\text{O(l)} + \text{KCl(aq)} \]

b. \[ \text{KIO}_3 + 2 \times \text{KI} + \text{H}_2\text{SO}_4 \rightarrow \text{K}_2\text{SO}_4 + \text{H}_2\text{O} + \text{I}_2 + \text{I}_2 \]

14. Use the oxidation number method to balance the following reactions. Show all work (oxidation numbers, which species are gaining and which are losing electrons, etc.)
a. \( \text{Mn}^{2+} + \text{BiO}_3^{3-} + \text{H}^+ \rightarrow \text{MnO}_4^{1-} + \text{Bi}^{3+} + \text{H}_2\text{O} \)

b. \( \text{HNO}_3(\text{aq}) + \text{H}_3\text{AsO}_3(\text{aq}) \rightarrow \text{NO}(\text{g}) + \text{H}_3\text{AsO}_4(\text{aq}) + \text{H}_2\text{O}(\text{l}) \)
Quiz 4.3 Redox

1. Use the oxidation number method to balance the following reactions. Show all work (oxidation numbers, which species are gaining and which are losing electrons, etc.)

   \[ \text{HNO}_3(aq) + \text{H}_2\text{S(aq)} \rightarrow \text{NO(g)} + \text{S(s)} + \text{H}_2\text{O(l)} \]

2. Which reactant is reduced during the reaction? (circle one) \( \text{HNO}_3(aq) \) or \( \text{H}_2\text{S(aq)} \)

3. Which reactant is the reducing agent? (circle one) \( \text{HNO}_3(aq) \) or \( \text{H}_2\text{S(aq)} \)