EXPERIMENT 15

Reactions of Elements - Composition, Replacement, and Combustion

INTRODUCTION

The chemical properties of elements and compounds are conveniently described in terms of the chemical reactions in which they participate. Although several thousand reactions have been studied, it is possible to group many of them as either oxidation-reduction reactions or non-oxidation-reduction reactions. When elements are the reactants in chemical reactions, they undergo oxidation-reduction reactions.

Oxidation-reduction, or redox, reactions feature a change in the oxidation states of two or more elements. Oxidation is defined as a process in which an element goes from a lower to a higher oxidation state. This happens when atoms of that element lose one or more electrons. Reduction involves a transition from a higher to lower oxidation state. This occurs when atoms of that element gain one or more electrons. An inherent feature of redox reactions is that both processes must occur simultaneously, that is, an element in one compound or ion can only experience an oxidation by causing the reduction of an element in another compound or ion. For this reason, the compound or ion containing the element undergoing oxidation is referred to as the reducing agent; the compound or ion containing the element undergoing reduction is called the oxidizing agent. This experiment will look at three types of redox reactions that elements undergo.

1. **COMPOSITION:** \( A + B \rightarrow AB \). Composition reactions are distinguished by a direct combination of elements or compounds to form a new compound. If the reactants are elemental forms of matter, the reaction must be an oxidation-reduction reaction because the neutral atoms of each element achieve either higher or lower oxidation states (lose or gain electrons) to form the product compound.

   In the above composition reaction, assume the element \( A \) forms stable 1+ ions and the element \( B \) forms stable 1- ions. In the reaction, neutral element \( A \) loses an electron to become a positive ion, while the neutral atom \( B \) gains the electron to become a negative ion. The oxidation half-reaction would be:
   \[
   A \rightarrow A^+ + e^-
   \]

   while the reduction half-reaction would be:
   \[
   B + e^- \rightarrow B^-
   \]

   When these two half-reactions are added together, the electron cancels out, and they yield:
   \[
   A + B \rightarrow A^+ + B^-
   \]

   or:
   \[
   A + B \rightarrow AB
   \]

   All oxidation-reduction reactions can be expressed as the sum of an oxidation half-reaction and a reduction half-reaction.
2. **REPLACEMENT:** \( A + BX \rightarrow AX + B \). Replacement reactions involve the transfer of electrons between atoms of an active element and the ions of a less active element. There are four common types of replacement reactions:

(a) An active elemental metal will react with dissolved ions of a less active metal. The active metal atoms lose electrons to form stable ions, while the less active metal ions gain the electrons to become elemental metal atoms. Assuming the metals each form stable 1+ ions and are bonded to anion \( X \) with a charge of 1-, the molecular equation for this reaction is:

\[
A(s) + BX(aq) \rightarrow AX(aq) + B(s)
\]

the ionic equation is:

\[
A(s) + B^+(aq) + X^-(aq) \rightarrow A^+(aq) + X^-(aq) + B(s)
\]

and canceling out the \( X^- \) spectator ions, the net ionic equation is:

\[
A(s) + B^+(aq) \rightarrow A^+(aq) + B(s)
\]

(b) An active elemental halogen will react with dissolved ions of a less active halogen. The active halogen atoms each gain an electron to form stable ions while the less active halide ions each lose an electron to become elemental halogen atoms. Because elemental halogens exist as diatomic molecules and form stable 1- ions, and assuming the halide ions are bonded to cation \( X \) with a charge of 1+, the molecular equation for this reaction is:

\[
A_2(g) + 2XB(aq) \rightarrow 2XA(aq) + B_2(g)
\]

the ionic equation is:

\[
A_2(g) + 2X^-(aq) + 2B^+(aq) \rightarrow 2X^+(aq) + 2A^+(aq) + B_2(g)
\]

and canceling out the \( 2X^- \) spectator ions, the net ionic equation is:

\[
A_2(g) + 2B^+(aq) \rightarrow 2A^+(aq) + B_2(g)
\]

(c) Active metals, which include almost all metals except the coinage metals, the noble metals, and mercury, will react with acids. The active metal atoms lose their valence electrons to form stable ions. The electrons are gained by hydrogen ions produced when the acid molecules ionize in solution. The hydrogen ions gain the electrons to become elemental hydrogen atoms. Reactions occur more readily with strong acids than with weak acids. Because hydrogen forms stable 1+ ions and in elemental form exist as diatomic molecules, and assuming that metal \( A \) forms stable 1+ ions and that the strong acid contains anion \( X \) with a charge of 1-, the molecular equation for this reaction is:

\[
2A(s) + 2HX(aq) \rightarrow 2AX(aq) + H_2(g)
\]

the ionic equation is:

\[
2A(s) + 2H^+(aq) + 2X^-(aq) \rightarrow 2A^+(aq) + 2X^-(aq) + H_2(g)
\]

and canceling out the \( 2X^- \) spectator ions, the net ionic equation is:

\[
2A(s) + 2H^+(aq) \rightarrow 2A^+(aq) + H_2(g)
\]
(d) Very active metals, such as the alkali and alkaline earth metals, will react with water. The very active metal atoms lose their valence electrons to form stable ions. The electrons are gained by hydrogen ions produced when water molecules slightly ionize into hydrogen ions and hydroxide ions. The hydrogen ions gain the electrons to become elemental hydrogen atoms. Because elemental hydrogen exists as diatomic molecules and forms stable 1+ ions, and assuming that metal $A$ forms stable 1+ ions, the molecular equation for this reaction is:

$$2A(s) + 2H_2O(l) \rightarrow 2AOH(aq \ or \ s) + H_2(g)$$

and this would be the ionic and net ionic equation if the metal hydroxide produced is insoluble. If the metal hydroxide produced is soluble, the ionic equation is:

$$2A(s) + 2H_2O(l) \rightarrow 2A^+(aq) + 2OH^-(aq) + H_2(g)$$

which is also the net ionic equation because there are no spectator ions.

3. **COMBUSTION.** Combustion reactions involve the reaction of either elements or compounds with elemental oxygen. Because at least one of the reactants is in its elemental form, this is evidence that the reaction must be an oxidation-reduction reaction. The reaction of an element with oxygen can also be classified as a composition reaction, which was covered earlier. In the combustion of a compound, the atoms of elemental oxygen achieve the lower oxidation state of -2 by gaining electrons from one of the elements in the compound, which achieves a higher oxidation state by losing electrons. Assuming that element $X$ maintains an oxidation state of +1, and element $A$ will change its oxidation state from -1 to +1, the molecular equation for this reaction is:

$$2XA + O_2 \rightarrow X_2O + A_2O$$

Notice that both $X$ and $A$ form compounds with the oxide ions produced. Because combustion reactions do not occur in water solution, they cannot be expressed in ionic form or net ionic form.

While classifying chemical reactions is often useful, it is more important to be familiar with the signs that indicate a chemical reaction is taking place. Evidence that a chemical change has occurred includes a color change, the production of light, or changes in energy. A temperature increase accompanies those reactions which evolve heat energy (called **exothermic** reactions) and a decrease in temperature is associated with those which absorb heat energy from their surroundings (called **endothermic** reactions). The evolution of a gas or the formation of a solid precipitate is also evidence that a reaction is occurring. While these signs are easily detected, others can only be perceived with the aid of instrumentation.

### PROCEDURE

1. Students will work individually for this experiment. Except for the laboratory handout, remove all books, purses, and such items from the laboratory bench top, and placed them in the storage area by the front door. For laboratory experiments you should be wearing closed-toe shoes. Tie back long hair, and do not wear long, dangling jewelry or clothes with loose and baggy sleeves. Open you lab locker. Put on your safety goggles, your lab coat, and gloves.

**REACTION 1 - COMPOSITION OF A METAL AND NONMETAL (DEMONSTRATION)**

2. Observe as the instructor heats a mixture of 2.0 grams of zinc and 1.0 gram of sulfur, and note the evidence of chemical action. Work out the balanced equation for the reaction between elemental zinc and elemental sulfur in the Reaction 1 box of the **Reaction Work** section following the Data Table. Write the balanced equation in the Reaction 1 box of the **Final Equations** section on page 155, and include the phases of each reactant and product in the equation. Record the following physical and chemical properties of elemental zinc, elemental sulfur, and the compound produced: color and reactivity with acid.
REACTION 2 - COMPOSITION OF A METAL AND NONMETAL

3. Obtain a piece of magnesium ribbon from the cart. Record the color and luster of the ribbon. Place a hot pad next to a burner. Holding one end of the magnesium ribbon with your crucible tongs, ignite the other end in the burner flame, and then hold the burning magnesium over the hot pad. Do not look directly at the magnesium while it is burning. Record the physical properties of the ash, which may be collected on the hot pad. Work out the balanced equation for the reaction in the Reaction 2 box of the Reaction Work section following the Data Table. Write the balanced equation in the Reaction 2 box of the Final Equations section on page 155, including the phases of each reactant and product.

REACTION 3 - COMPOSITION OF TWO NONMETALS

4. Pour about 50 mL of deionized water into a 250-mL beaker. Position the snorkel hood so that it is about 15 inches above the lab bench, and set the beaker under the snorkel hood. Obtain a deflagrating spoon from the cart and half fill it with elemental sulfur. With your burner also placed under the fume hood at our lab station, heat the sulfur with the burner flame until it ignites with a pale blue flame. CAUTION: The pale blue sulfur flame is hard to see and can cause severe burns.

Hold the deflagrating spoon under the snorkel hood and let the sulfur completely burn away. When the sulfur has all burned away, carefully lower the deflagrating spoon into the beaker of water to cool it off. The deflagrating spoon should now be clean and free of any black sulfur. Dry the deflagrating spoon and return it to the cart. Record the physical properties of the reaction product. The sulfur burning in oxygen produces a product in which the oxidation state of sulfur is +4. Work out the balanced equation in the Reaction 3 box of the Reaction Work section following the Data Table. Write the balanced equation in the Reaction 3 box of the Final Equations section on page 155, including the phases of each reactant and product.
REACTION 4 - METALS REPLACING HYDROGEN IN WATER

5. Take a watch glass to the desiccator in Fume Hood A, and using the forceps provided, place a small piece of calcium metal on the watch glass and return it to your lab station.

   **CAUTION:** Do not touch calcium metal, and handle with care.

   Place about 3 mL of deionized water into three medium test tubes. Measure the 3 mL with a graduated cylinder one time only, then approximate all further volumes. With the test tubes supported in your test tube rack, add a small piece of copper metal to the first, a small piece of zinc metal to the second, and a small piece of calcium metal to the third.

   **NOTE:** In Reation 4, a slight color change to the solution is not evidence of the metal reacting with the water, it is actually an oxide coating on the metal reacting, so it should be ignored.

   Record your observations as to whether chemical reactions have occurred or not.

6. For test tubes in which reactions occurred, test their contents with litmus paper. Obtain a strip of red litmus paper (it looks pink). Using your clean, dry stirring rod, dip the stirring rod into the solution, and then touch the stirring rod to the litmus paper. Red litmus paper turning blue (it looks purple) shows that the solution tested is basic. This indicates that a product of the chemical reaction is a base. Record the color the solution turns the litmus paper in your Data Table. If a reaction occurred, work out the balanced molecular and ionic equations in the Reaction 4 box of the Reaction Work section following the Data Table. Write the balanced net ionic equation in the appropriate Reaction 4 box of the Final Equations section on page 155 (4A for copper, 4B for zinc, and 4C for calcium). If no reaction occurred, write NR in the Final Equations section for the specific reaction. Dispose of these solutions down the sink, and save them for Reaction 5.

REACTION 5 - METALS REPLACING HYDROGEN IN ACIDS

7. Place 3 mL of 6 M hydrochloric acid into two medium test tubes.

   **CAUTION:** Hydrochloric acid is corrosive and can cause severe burns.

   Add a small piece of copper metal to the first and a small piece of zinc metal to the second.

   **NOTE:** In Reation 5, a color change to the solution is not evidence of the metal reacting with the hydrochloric acid, it is actually an oxide coating on the metal reacting, so it should be ignored.

   Record your observations as to whether chemical reactions have occurred or not.

8. If a gas is produced, hold your thumb over the mouth of the test tube for about one minute, light a wooden splint with a burner, then remove your thumb and hold the burning splint to the mouth of the test tube. Record your observations and identify the gas. If a reaction occurred, work out the balanced molecular and ionic equations in the Reaction 5 box of the Reaction Work section following the Data Table. Write the balanced net ionic equation in the appropriate Reaction 5 box of the Final Equations section on page 155 (5A for copper and 5B for zinc). If no reaction occurred, write NR in the Final Equations section for the specific reaction. Dispose of these solutions down the sink, and save the unreacted metals for Reaction 6.
REACTION 6 - METALS REPLACING METAL IONS

9. Place 3 mL of zinc sulfate solution in a small test tube. Add a piece of copper metal and observe the test tubes for several minutes. Record your observations as to whether a chemical reaction has occurred or not. If a reaction occurred, work out the balanced molecular and ionic equations in the Reaction 6 box of the Reaction Work section following the Data Table, and write the balanced net ionic equation in the Reaction 6A box of the Final Equations section on page 151. If no reaction occurred, write NR in the Reaction 6A box of the Final Equations section.

10. Place 3 mL of copper (II) sulfate solution in a small test tube. Add a piece of zinc metal. Mix with a stirring rod once each minute while observing the test tube for ten minutes. Record your observations as to whether a chemical reaction has occurred or not. If a reaction occurred, work out the balanced molecular and ionic equations in the Reaction 6 box of the Reaction Work section following the Data Table, and write the balanced net ionic equation in the Reaction 6B box of the Final Equations section on page 155. If no reaction occurred, write NR in the Reaction 6B box of the Final Equations section. Place these solutions in a waste beaker on your lab bench and dispose of them in the Liquid Waste bottle in the fume hood. The metals can be disposed of in the trashcan.

REACTION 7 – COMBUSTION REACTION

11. Light your laboratory burner to burn the natural gas (methane) coming from the gas outlet at your lab station. Remove any flammable material from your lab bench. Holding the burner by its base, turn the burner over and have the flame come in contact with the lab bench for 2 or 3 seconds. Place the burner back on the lab bench and turn it off. Record your observations of the product seen on the lab bench, and identify it. Work out the balanced equation for the burning of methane in the Reaction 7 box of the Reaction Work section following the Data Table. Write the balanced equation in the Reaction 7 box of the Final Equations section on page 155.

12. Clean and wipe dry your laboratory work area and all apparatus. When you have completed your lab report have the instructor inspect your working area. Once your working area has been checked your lab report can then be turned in to the instructor.
## EXPERIMENT 15 LAB REPORT

Name: ________________________________  Student Lab Score: ______________

Date/Lab Start Time: ____________________  Lab Station Number: ______________

### DATA TABLE

<table>
<thead>
<tr>
<th>REACTION 1 OBSERVATIONS</th>
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<tbody>
<tr>
<td>Zinc and Sulfur with Heat</td>
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<table>
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<tr>
<th>PHYS/CHEM PROPERTIES</th>
<th>SULFUR</th>
<th>ZINC</th>
<th>PRODUCT</th>
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<tbody>
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<tr>
<td>Substance with Acid</td>
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<th>GASEOUS PRODUCT OBSERVATIONS</th>
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<td>Identifying Characteristic of Gaseous Product</td>
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<td>Identity of Gaseous Product</td>
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<th>REACTION 2 OBSERVATIONS</th>
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<td>Magnesium Ribbon with Heat</td>
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<tr>
<th>REACTION 3 OBSERVATIONS</th>
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<td>Sulfur with Heat</td>
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### REACTION 4 OBSERVATIONS

- Copper with Water
- Zinc with Water
- Calcium with Water
- Litmus Paper with Product Solution

### REACTION 5 OBSERVATIONS

- Copper with Hydrochloric Acid
- Zinc with Hydrochloric Acid
- Observation of Gaseous Product with Burning Splint
- Identity of Gaseous Product

### REACTION 6 OBSERVATIONS

- Zinc Sulfate with Copper
- Copper (II) Sulfate with Zinc

### REACTION 7 OBSERVATIONS

- Observation of Methane with Spark
- Observation of Reaction on the Lab Bench
- Identity of Product on the Lab Bench
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**FINAL EQUATIONS**

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<td>Reaction 6B</td>
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<tr>
<td>Reaction 7</td>
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</table>
1. Write the oxidation half-reaction and the reduction half-reaction for the reaction between zinc and sulfur in Reaction 1.

2. Write the oxidation half-reaction and the reduction half-reaction for the reaction between zinc and hydrochloric acid in Reaction 5.

3. Write the oxidation half-reaction and the reduction half-reaction for the reaction between zinc and copper (II) sulfate in Reaction 6.