Lesson Plan: Single Displacement Reactions with Test Tube Diagrams

FOR THE TEACHER

Summary
In this lesson students will perform and analyze two single displacement reactions and prepare and manipulate Test Tube Diagrams to depict the activity at the molecular level. Using manipulatives representing individual ions, atoms and molecules for the various reactants and products, they will accurately represent species in the solid, gaseous and aqueous states by correlating the Test Tube Diagram to the complete ionic equation for each reaction. They will determine the reactants and products responsible for color, as well as identify which species is oxidized and which is reduced.

Grade Level
High School

NGSS Alignment
This lesson will help prepare your students to meet the performance expectations in the following standards:

- **HS-PS1-5**: Apply scientific principles and evidence to provide an explanation about the effects of changing the temperature or concentration of the reacting particles on the rate at which a reaction occurs.
- **HS-PS1-7**: Use mathematical representation to support the claim that atoms, and therefore mass, are conserved during a chemical reaction.
- **HS-PS2-6**: Communicate scientific and technical information about why the molecular-level structure is important in the functioning of designed materials.
- **Scientific and Engineering Practices**:
  - Using Mathematics and Computational Thinking
  - Developing and Using Models
  - Analyzing and Interpreting Data

Objectives
By the end of this lesson, students should be able to

- Determine whether a single displacement reaction will occur.
- Accurately represent the activity of a single displacement reaction occurring at the molecular level.
- Identify the species that is oxidized and the species that is reduced.
- Demonstrate the Law of Conservation of Mass using manipulatives and relate it to a balanced chemical equation.
- Depict the progress of a reaction using postulates of the Kinetic Molecular Theory.

Chemistry Topics
This lesson supports students’ understanding of

- Reactions
- Single Displacement Reactions
- Activity Series
- Law of Conservation of Mass
- Redox Reactions
- Oxidation
• Reduction
• Net Ionic Equations
• Kinetic Molecular Theory
• Evidence of a Chemical Reaction

Time
Teacher Preparation: 30 minutes – 2 hours (dependent on need to create manipulatives)

Lesson:
• Engage: 5-10 minutes
• Explore: 60 - 100 min
• Explain: 20-30 minutes
• Elaborate: 30 – 60 minutes
• Evaluate: 20-30 minutes

Note: Ideally Explore through Evaluate are happening simultaneously. The writing prompt (Elaborate) may be completed after the activity.

Materials
(Engage)
• 10 cm piece of copper wire, sanded (wire coiled around a pencil is more interesting visually.)
• 10 mL of 0.1 M silver nitrate solution
• 1 gas collection bottle of oxygen
• 20 cm rolled piece of steel wool
• Tongs
• Bunsen burner
• Glass plate
• Gas collection trough
• 1 cm x 10 cm Zn strip
• 0.2 M CuSO₄
• 10 cm Cu wire
• 0.2 M ZnSO₄

(Explore & Explain)
Per lab group:
• 2, 5cm strips of Magnesium ribbon
• 2, 1cm x 5cm Aluminum strips/foil
• 15 mL of 1 M Nitric acid
• 30 mL of 0.2 M Copper (II) chloride solution
• Before & After Test Tube Diagram sheets (cardstock and/or laminated)
• Various colors of cardstock/paper for cut-outs:
  o Silver or gray (Al & Mg)
  o Brown (Cu)
  o Blue (Cu²⁺)
  o White (Al³⁺, H⁺, Mg²⁺, NO₃⁻, Cl⁻, and H₂)
• Student Kit Contents Procedure 1 (see teacher notes for photo example)
  o 4 silver Mg atom cards, 4 white Mg²⁺ ion cards, 8 white H⁺ ion cards, 4 white H₂ molecule cards, and 8 white NO₃⁻ ion cards.
• Student Kit Contents Procedure 2 (see teacher notes for photo example)
  o 4 silver Al atom cards, 4 white Al³⁺ ion cards, 8 brown Cu atom cards, 8 blue Cu²⁺ ion cards, and 16 white Cl⁻ ion cards.
Safety

- Students should wear proper safety gear during chemistry demonstrations. Safety goggles and lab apron are required. Gloves may also be worn.
- Procedure 1 produces hydrogen gas. If a burning splint is brought to the mouth of the test tube during the first several minutes of production, it may be possible to ignite the hydrogen gas. There should be no open flames at the benches during Procedure 1. Magnesium is the limiting reactant, approximately 0.07 g, and is capable of producing only 0.13 L of hydrogen. Twelve lab pairs will produce only 1.6 L of hydrogen gas over the five to ten minutes required for the reaction to go to completion.
- Always wear safety goggles when handling chemicals in the lab.
- When working with acids, if any solution gets on students’ skin, they should immediately alert you and thoroughly flush their skin with water.
- Procedure 2 requires a hot plate or a burner.
  - Always use caution around open flames. Keep flames away from flammable substances.
  - Always be aware of an open flame. Do not reach over it. Tie back hair, and secure loose clothing.
  - Open flames can cause burns.
  - Exercise caution when using a heat source. Hot plates should be turned off and unplugged as soon as they are no longer needed.
  - When lighting a match, be cautious with the flame.
  - An operational fire extinguisher should be in the classroom.
- When students complete the lab, instruct them how to clean up their materials and dispose of any chemicals.
- Students should wash their hands thoroughly before leaving the lab.
- For more chemical information refer to this link.
- For disposal information refer to this link.

Teacher Notes

- **Preparation:** Like any other chemistry lab, metal and solution preparation will require 15 – 30 minutes depending on whether the teacher cuts the Mg and Al into strips or has students cut them.
- The atom/molecule/ion cards for each lab pair could require a couple hours to prepare the first time if you do all the cutting yourself. Make extra sheets while initially copying on the appropriately colored cardstock or paper to replace lost cards in the student kits. Zip-Lock bags or baby food jars can be used for the lab pair sets. Numbering them may make students feel more accountable for maintaining the inventory in each student set. Keep Procedure 1 and Procedure 2 cards in separate containers.
- Each class can make sure the correct number of each card is in the kit and replace missing cards. Have the extra cards in Ziplocs at the ready for this. The kit for Procedure 2 only contains 40 cards and is still a quick count for two people, especially since there are several colors and both sides of the cards have the formula. Each kit has several extra cards for each species, just in case. This also allows the kits to be used to illustrate limiting/excess reactants during a unit on stoichiometry. Once the student atom/molecule/ion card sets are originally prepared, they can be used year after year.
- For Procedure 1 you will need one discard beaker labeled Discard Procedure 1: Mg + HNO₃. This beaker should contain only Mg(NO₃)₂, the excess HNO₃, and perhaps some Mg (unlikely).
- For Procedure 2 you will need three discard beakers labeled 1) Discard Procedure 2 Reaction: Al + CuCl₂, 2) Discard Procedure 2: Reference CuCl₂, and Discard Procedure 2: Reference Al.

- **Engage:** Although the focus here is on single displacement reactions, the thinking can and should be extended more generally to any oxidation/reduction process. Three quick demos will catch their attention.
Demo 1: The classic copper wire in a solution of silver nitrate can be used. Write the following equation on the board with the help of the students and the Activity Series.

$$
\text{Cu (s)} + \text{AgNO}_3 (\text{aq}) \rightarrow \text{Cu(NO}_3\text{)}_2 (\text{aq}) + \text{Ag (s)}
$$

**Procedure:**
1. Dip the sanded Cu wire in and *immediately* take it out.
2. The color change on the copper wire indicates just how quickly the reaction occurs.
3. Replace it in the solution and pass it around the class.
4. Warn students that silver nitrate solution will stain the skin.
5. Place a stopper on the test tube before passing it around if you wish. Have students pass it quickly the first time around. Then pass it around again so they can inspect it more thoroughly. Students are amazed that the “fuzz, mold” growing on the clean copper wire is actually pure silver.
6. Considering that the chemical equation is on the board, their inability to correlate the chemical equation (the abstract) with their observations (the concrete) is apparent. One of the goals of this activity is to address these gaps in knowledge and to correlate students’ observations of the chemical reaction with the written chemical equation.

Demo 2: Similarly, write the equation for the reaction between Zn and CuSO₄ solution on the board.

**Procedure:**
1. Dip the zinc strip into the CuSO₄ solution. Students will observe an immediate reaction.
2. Consult the Activity Series.
3. Now write the equation for the reaction between Cu and ZnSO₄ on the board.
4. Dip the Cu wire into the ZnSO₄ solution. Students will observe no reaction.
5. Consult the Activity Series.

Demo 3: Procedure:
1. Using a gas collection trough, prepare a bottle of oxygen by the catalytic (MnO₂, KI or yeast) decomposition of 6% hydrogen peroxide.
2. Ignite a 20 cm long rolled piece of course steel wool in a burner, quickly place it into the bottle of oxygen, and observe the fireworks.
3. Metals love oxygen and oxygen loves metals.

**Explore:** This is where students observe, at the macroscopic level, two single displacement reactions and try to reconcile their observations with the molecular level activity responsible for the changes they witness. They will perform each reaction, write molecular and ionic equations for each reaction and prepare Test Tube Diagrams. The instructor can view the diagrams and quickly ascertain student understanding of the difference between a neutral atom and its aqueous ion, and a
solid, gaseous or aqueous species, what species are responsible for the colors observed, electron activity and collision theory.

- Students can predict products when given reactants and balance chemical equations without having any real sense of the activity at the molecular level, without understanding the chemistry occurring.
- Ask students where the piece of aluminum is in the test tube, or the hydrogen or the copper ion. Ask them how they know. Ask them if they have realistically shown each of those in their Test Tube Diagrams.
- Were copper atoms or aluminum atoms created or destroyed? Were more electrons lost than gained? What do their diagrams say about these issues?
- Ask them to move a hydrogen ion around until it is in a position to accept an electron from a magnesium atom.
- Help the students focus on the main ideas – atoms have to collide in order to react; atoms, compounds and solutions are electrically neutral; atoms are conserved; oxidation/reduction are components of a single process; some species is responsible for the color we see, or none of the species yield a color; and that we need to somehow protect metals from reacting when we use them to construct bridges, buildings, cars, tableware ...
- Prior to this lab I have already introduced students to general reaction types including single displacement reactions. They are familiar with the Activity Series and how it is used. They have predicted whether a reaction will occur and have seen examples that did not yield a reaction, copper in zinc sulfate or my wedding ring in hydrochloric acid, for example.
- Students have also hand drawn test tube diagrams representing the species present in before and after test tubes. However, as this lab activity demonstrates, they have very limited understanding of the abstraction that is a balanced complete ionic equation. As they play with the manipulatives at each bench there are numerous “Ah-Hah!” moments followed by “That makes sense!” comments. Physical states, oxidation states, collision theory all come to light as they physically run and observe the progress of the reactions while simultaneously and accurately creating diagrams to account for the chemistry they are witnessing.

- **Student lab Procedure #1**: Single Displacement / Test Tube Diagram Manipulatives:
  
  - These Test Tube Diagrams show the ions in solution well organized, nice to read but unrealistic. The Mg and H$_2$ are realistically placed. Note Mg (s) is on the bottom and H$_2$ (g) is up and out of the test tube. H$^{1+}$ (aq) and NO$_3^{-}$ (aq) are floating in the solution. Also note that the Mg (s) has been completely converted to Mg$^{2+}$ in the After Test Tube. It was the limiting reactant. If you would like to show that HNO$_3$ is in excess, add another molecule or two of it to both the Before and After test tubes. For this lab have students place stoichiometric amounts, as indicated in the balanced chemical equation (1 Mg & 2 HNO$_3$; 1 Mg$^{2+}$ & 2 NO$_3^{-}$ & 1 H$_2$), in the Before & After test tubes.
  
  - These Test Tube Diagrams show ions more randomly placed, more realistically. Note Mg (s) is on the bottom, H$_2$ (g) is up and out of the test tube and Mg$^{2+}$ (aq) and NO$_3^{-}$ (aq) are floating in the solution.
- This Before Diagram has the H\(^{+}\) ions striking the Mg (s) providing for the opportunity for electron exchange leading to the formation of H\(_2\) (g) and Mg\(^{2+}\). Students can randomly move the H\(^{+}\) ions around in the beaker until they strike the Mg emphasizing that reaction occurs when these species collide. The Mg atoms (solid) cannot move about in the beaker. However, the aqueous H\(^{+}\) ions are free to move throughout the solution and will eventually bump into the Mg ribbon. In practice, students should note that this occurs instantaneously and continuously. And that collisions become less likely as the concentration of the H\(^{+}\) decreases.

- **Student lab Procedure #2: Single Displacement / Test Tube Diagram Manipulatives:**

- These Test Tube Diagrams show ions well organized, nice to read but unrealistic. For problem set, quiz and test questions, this may be a good way to have students draw them for ease of counting for both the student and the teacher. Note Al (s) & Cu (s) are on the bottom. Cu\(^{2+}\) (aq), Cl\(^{-}\) (aq) and Al\(^{3+}\) are floating in the solution. Also note that the Al (s) is completely gone. This time, in reality, the solid is not the limiting reactant, the metal ions in solution are limiting. CuCl\(_2\) is the limiting reactant. If you would like to show that Al is in excess, add another atom or two of it to both the Before and After test tubes. Since Al remains after reaction ceases and since the Cu\(^{2+}\) ions are responsible for the blue color of the solution, which has completely disappeared after 5 or 10 minutes, the Cu\(^{2+}\) ions (CuCl\(_2\)) must be limiting.

- These Test Tube Diagrams show ions more randomly placed, more realistically. Note Al(s) & Cu (s) are on the bottom. Cu\(^{2+}\) (aq), Cl\(^{-}\) (aq) and Al\(^{3+}\) (aq) are floating in the solution. The teacher can evaluate this aspect of the student Test Tube Diagrams as he/she walks about the lab checking and discussing student diagrams. At the end of the class, the teacher can highlight or have students highlight all of the main points while the diagrams are still intact.

- This Before Diagram has the Cu\(^{2+}\) ions striking the Al\(^{0}\) (s) providing for the opportunity for electron exchange leading to the formation of Al\(^{3+}\) (aq) and Cu\(^{0}\) (s). Students can randomly move the Cu\(^{2+}\) ions around in the beaker until they strike the Al emphasizing that reaction occurs when these species collide. The Al atoms (solid) cannot move about in the beaker. However, the aqueous Cu\(^{2+}\) ions are free to move throughout the beaker and will eventually bump into the Al strip. In practice, students should note that this occurs instantaneously and continuously. As this
reaction progresses students can witness the disappearance of the blue color as the concentration of Cu²⁺ decreases.

- **Explain:** I find it convenient to have students artificially line the ions up on the diagrams. As you walk around the room you can quickly count ions this way and then randomize the diagram yourself, or start to and have the students finish randomizing it. You can ask the students questions about the locations, changes in colors, charges, etc. of the various species. Use the student handout to guide you in asking questions of the students. If a misconception arises consistently, you can address the entire class at once and have them adjust their diagrams to make a point or to clarify something. You can emphasize the balance of charge that has to exist in both Before and After test tubes. You can use the diagrams to illustrate the conservation of atoms and of mass. All of these concepts are addressed on their lab reports. Have them complete the student questions while they have the diagrams to refer to and to manipulate. Remember the manipulatives have the colored species, the students’ hand-written reports will not. If you wish, have the students color their hand-written reports: gray or silver (Mg & Al), brown (Cu), and blue (Cu²⁺).

- **Extend/Elaborate:** Questions on the lab report address the concept of oxidation/reduction more generally.
  - Students poured some of the solution from Procedure 2 onto a watch glass. Question 16 asks them what they expect to find on the watch glass once the water evaporates. Remember to have students check the watch glasses in a few days.
  - *How to prevent Metals from Corroding* is a good resource for the “how metals can be protected” research for students.
  - The Writing prompts give a new scenario for students to evaluate in terms of particle locations, charges, oxidation/reduction and colors. This is intended to be given a day or two after this activity. See *Collins Writing Approach* for more information.

- **Additional Information:** These kits can also be used when you discuss stoichiometry. Use them as above to demonstrate stoichiometric amounts. Add one extra formula for the excess reactant than the number called for stoichiometrically. For example, if 50.0 g of Mg reacts with 50.0 g of HNO₃, then Mg is in excess. Place a second Mg card in the Before diagram – the balanced equation calls for one Mg. In the After diagram, place one Mg²⁺ card to represent the magnesium that reacted and also one Mg card representing the excess Mg that did not react. If 5.00 g of Mg reacts with 50.0 g of HNO₃, then HNO₃ is in excess. Place three H⁺ cards and three NO₃⁻ cards in the Before diagram. In the After diagram, place one H₂ card, one H⁺ card and one NO₃⁻ card. You can show the H⁺ moving about in the After diagram “looking” for a Mg to react with. But there is no Mg, only Mg²⁺.

- See photo below for Student Kit Contents for Reaction #1: Mg + HNO₃
- See photo below for Student Kit Contents for Reaction #2: Al + CuCl₂
FOR THE STUDENT
Lesson

Single Displacement Reactions Including Test Tube Diagrams

Background
Most elements are quite reactive and, as a result, are not naturally found in their pure elemental form but are found in compounds. This is certainly true of most metals and for hydrogen. Most metals react with nonmetals and are oxidized, losing electron(s), in the process. The nonmetals take the electron(s) and are reduced. Consider the reaction between zinc and sulfur producing zinc sulfide.

\[ \text{Zn}^0 (s) + \text{S}^0 (s) \rightarrow \text{Zn}^{2+}\text{S}^{2-} (s) \]

Zinc has lost two electrons to sulfur, so zinc has been oxidized. Zinc’s charge or oxidation state has changed from 0 to 2+. Sulfur has gained two electrons from zinc, so sulfur has been reduced. Sulfur’s charge or oxidation state has changed from 0 to 2⁻.

Metals may be present in the neutral, elemental state, such as the reactant zinc in the above equation. Or they may be present as a cation, such as zinc in the product ZnS in the above equation. In compounds, metals are present in the cation form.

In addition to reacting with nonmetals by transferring electrons to them, neutral metals may also be able to transfer electrons to the cation form of another metal. Metals have a “pecking” order, which is called an Activity Series. Free metals are able to transfer electrons to the metal cation in a compound if they are more active than the metal in the compound. When this occurs, the free metal is said to have displaced the metal in the compound.

A metal’s activity is really a measure of its tendency to give up electrons during a chemical reaction relative to that of other metals. Refer to the Activity Series in your textbook. The metals at the top are the most likely give up an electron(s), changing from the neutral atom to the positive ion by transferring electrons to the cation of the other metal. That is the metals at the top of the Activity Series are most readily oxidized. The cation, the form of the metal that is lower on the chart, receives the electrons and is reduced changing from a positive ion to a neutral atom. These processes occur most easily when the compound is dissolved in water so the cation and anion are free to move independently of each other. When the random motion of the metal cation in the solution causes it to collide with a neutral metal atom of the free element there may be an exchange of electrons from the neutral metal atom to the metal cation. This exchange occurs only if the free element has a greater tendency to form its cation than the metal already existing in the ionic state. This tendency is reflected in the Activity Series of metals.

For example, zinc is above copper on the Activity Series indicating that elemental zinc loses electrons more readily than elemental copper. Therefore, when a copper(II) ion, \( \text{Cu}^{2+} \), in solution collides with an atom of zinc, \( \text{Zn}^0 \), two electrons will be transferred to the copper(II) ion from the zinc atom. This results in \( \text{Zn}^{2+} \) and \( \text{Cu}^0 \). As an ion, the zinc is now dissolved in the water and the now neutral copper crystallizes out of solution. Consider the reaction between a strip of zinc metal and a solution of copper(II) sulfate.
The molecular equation representing this process is

\[ \text{Zn (s)} + \text{CuSO}_4 \text{(aq)} \rightarrow \text{ZnSO}_4 \text{(aq)} + \text{Cu (s)} \]

The complete ionic equation representing this process is

\[ \text{Zn}^0 \text{(s)} + \text{Cu}^{2+} \text{(aq)} + \text{SO}_4^{2-} \text{(aq)} \rightarrow \text{Zn}^{2+} \text{(aq)} + \text{SO}_4^{2-} \text{(aq)} + \text{Cu}^0 \text{(s)} \]

The net ionic equation for this process is

\[ \text{Zn}^0 \text{(s)} + \text{Cu}^{2+} \text{(aq)} \rightarrow \text{Zn}^{2+} \text{(aq)} + \text{Cu}^0 \text{(s)} \]

Look at the Activity Series in your text book. Notice that hydrogen is included among the metals. Although not a metal, hydrogen does play the role of the “cation” in covalent molecular acids such as HCl and HNO\(_3\) where it has, not a charge of 1+, but an oxidation state of 1+. However, when dissolved in water acids ionize generating the hydrogen ion, H\(^{1+}\), and the anion of the acid, Cl\(^{-}\) or NO\(_3^{-}\), for example. The H\(^{1+}\) ion will be reduced to H\(_2^0\) when a metal above it, a metal with a greater tendency to lose an electron(s) than hydrogen, is placed in an acidic solution.

If the metallic element is below the metal or hydrogen in the compound, then no electrons will be transferred. In such a case, the free neutral metal is not capable of transferring its valence electron(s) to the cation, or hydrogen, and no reaction occurs. For example, if you drop your gold ring into a beaker containing hydrochloric acid, nothing will happen.

\[ \text{Au (s)} + \text{HCl (aq)} \rightarrow \text{No Reaction} \]

When a transfer of electrons occurs, the atom releasing the electrons undergoes an increase in its oxidation state from zero to a positive value equal to the number of electrons lost. In our example, Zn loses two electrons and changes from Zn\(^0\) to Zn\(^{2+}\). Zinc has been oxidized. The cation accepting the electrons undergoes a decrease in oxidation state from some positive value to zero. The copper(II) ion takes the two electrons from the zinc atom and changes from Cu\(^{2+}\) to Cu\(^0\). Copper has been reduced.

<table>
<thead>
<tr>
<th>Oxidation</th>
<th>3+</th>
<th>2+</th>
<th>1+</th>
<th>0</th>
</tr>
</thead>
<tbody>
<tr>
<td>Loss of Electron(s)</td>
<td>2-</td>
<td>1-</td>
<td>0-</td>
<td>1+</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Reduction</th>
<th>Gain of Electron(s)</th>
<th>1-</th>
<th>2-</th>
<th>3-</th>
</tr>
</thead>
<tbody>
<tr>
<td>Charge or Oxidation State</td>
<td></td>
<td>1-</td>
<td>2-</td>
<td>3-</td>
</tr>
</tbody>
</table>

In this example, zinc and copper react in a one to one mole ratio because zinc loses two electrons and copper gains two electrons. If the number of electrons lost by an atom of the free element does not equal the number of electrons gained by a single ion in solution, then they will not react in a one to one ratio. So, for example, when Al\(^0\) reacts with H\(^{1+}\) from HCl, three hydrogen ions would be required to accept the three electrons an aluminum atom will release. They react in a 3:1 ratio, H\(^{1+}\): Al\(^0\).
Consider the zinc and copper(II) sulfate reaction again. The complete ionic equation for this reaction is:

\[
\text{Zn}^0 (s) + \text{Cu}^{2+} (aq) + \text{SO}_4^{2-} (aq) \rightarrow \text{Zn}^{2+} (aq) + \text{SO}_4^{2-} (aq) + \text{Cu}^0 (s)
\]

- The zinc has undergone an oxidation changing from Zn\(^0\) (s) to Zn\(^{2+}\) (aq).
- The copper has undergone a reduction changing from Cu\(^{2+}\) (aq) to Cu\(^0\) (s).
- Zinc released two electrons, copper(II) took two electrons.
- The reaction occurred between zinc and copper(II).
- The sulfate ion began the reaction as SO\(_4^{2-}\) (aq); it ended the reaction as SO\(_4^{2-}\) (aq). Sulfate underwent no change. Sulfate did not react. Sulfate is called a spectator ion. It is necessary because there cannot be a cation (Cu\(^{2+}\)) without a compensating anion (SO\(_4^{2-}\)); compounds are electrically neutral.

**Pre-lab Questions**

1. Consider the following combinations. Using the Activity Series, decide whether or not a reaction occurs. To the right of the arrow, write “Rxn Occurs” or “No Rxn.”
   a. \( \text{Zn} (s) + \text{H}_3\text{PO}_4 (aq) \rightarrow \)
   b. \( \text{Ag} (s) + \text{Ca(NO}_3\text{)} (aq) \rightarrow \)
   c. \( \text{Ni} (s) + \text{Au(NO}_3\text{)}_3 (aq) \rightarrow \)
   d. \( \text{Cu} (s) + \text{AlCl}_3 (aq) \rightarrow \)

2. List the seven diatomic elements:

3. Consider the following combinations. Using the Activity Series, decide if the reaction occurs and if so, complete and balance the molecular equation.
   a. \( \text{Al} (s) + \text{HCl} (aq) \rightarrow \)
   b. \( \text{Mg} (s) + \text{Fe(NO}_3\text{)}_3 (aq) \rightarrow \)
   c. \( \text{Cu} (s) + \text{SnSO}_4 (aq) \rightarrow \)

**Objective**

- You will use the Activity Series to determine whether a reaction will occur.
- You will express the reactions using balanced molecular, complete ionic and net ionic equations including states.
- You will realistically display the reaction occurring in the test tube using cards that represent the atoms, ions and molecules involved in the reaction. To do this, you will consider physical states, density, the Law of Conservation of Mass and the conservation of charge.
- You will determine what atom or ion is responsible for any color, which species has been oxidized and which species has been reduced, and which ion is the spectator ion.
- Using Kinetic Molecular Theory, you will demonstrate/describe the process occurring in solution resulting in the transfer of electrons.
Materials

- 18 x 150 mm test tubes (5)
- 250 mL beaker
- 5 cm strips of magnesium ribbon (2)
- 1 cm x 5 cm strips of aluminum or Foil (2)
- 15 mL of 1.0 M nitric acid
- 30 mL of 0.2 M copper(II) chloride
- Hot plate (or 1 burner, ring stand & ring)
- Thermometer
- Watch Glass
- Wire gauze
- Full sheet of 80 grit sandpaper (steel wool)
- Set of Manipulatives for each procedure

Safety

- Always wear an apron and safety goggles when handling chemicals in the lab. Gloves may also be worn.
- Wash your hands thoroughly before leaving the lab.
- Follow the teacher’s instructions for cleanup of materials and disposal of chemicals.
- When working with acids and bases, if any solution gets on your skin immediately rinse the area with water.
- Always use caution around open flames. Keep flames away from flammable substances.
- If a burner is used:
  - Always be aware of an open flame. Do not reach over it. Tie back hair, and secure loose clothing.
  - When lighting a match, be cautious with the flame.
  - Open flames can cause burns.
- If a hot plate is used
  - Exercise caution when using a heat source. Hot plates should be turned off and unplugged as soon as they are no longer needed.

Procedure 1

Reaction between Magnesium (s) and Nitric Acid (aq)

SAFETY NOTE: There are to be no open flames or heat sources at your bench during Procedure 1.

1. Obtain 15 mL of 1.0 M nitric acid and pour it into a clean test tube labeled with “Acid.” Place 15 mL of room temperature tap water in a second test tube labeled “Water.” This test tube will serve as a reference for the initial appearance of your reaction test tube, the “Acid” test tube. It will help you determine the changes that are occurring in the reaction test tube over time. Record the temperature of the liquids in both test tubes in the data table.
2. If instructed to do so, lightly sand two 5 cm pieces of magnesium ribbon. Be careful, the edges are sharp.
3. Describe the appearance of the nitric acid, the strip of magnesium, and the water in the data table.
4. Drop one magnesium strip into the “Water” test tube. Push the strip down with a stirring rod if necessary. Drop the second magnesium strip into the “Acid” test tube. Push the strip down with a stirring rod if necessary.
5. Make initial observations of the reaction and record your observations in the data table. Make additional observations after reaction has been proceeding for 2 and 5 minutes. Feel both test tubes when making the 5 minute observations. Also take the
temperature of both test tubes at the 5 minute observation.
6. Write the molecular, complete ionic, and net ionic equations for the reaction of magnesium with nitric acid in the Analysis section.
7. Based on the complete ionic equation, use the element and ion cards to complete the Test Tube Diagrams on the Before and After Test Tube sheets. Place all cards in realistic locations (solid, gas, aqueous) on the Before and After Test Tube sheets.
8. Have the instructor check your Before and After Test Tube sheets.
9. After receiving instructor approval, complete the Test Tube Diagrams in the Analysis section and answer the questions for Procedure 1. Also answer questions pertaining to Procedure 1 found in the Conclusions and Extensions section.
10. After 10 minutes, make observations again.
11. If at this point the magnesium strip is completely dissolved, discard the contents of the test tube in the sink with running water. If the reaction is not yet complete, some magnesium remains, place the test tube upright in a beaker and move it to the side of the lab bench opposite the location of the hotplate or burner.

Procedure 2
The Reaction between Aluminum (s) and Copper(II) Chloride (aq)

1. Place 150 mL of tap water in a 250 mL beaker and start a hot water bath.
2. Obtain 15 mL of 0.2 M copper (II) chloride and pour it into a clean test tube labeled “Rxn.” Pour a second 15 mL sample of 0.2 M copper (II) chloride into a test tube labeled “Ref CuCl2.” Pour 15 mL of tap water into a third test tube and label it “Ref Al.” The Ref (reference) test tubes will serve as references for the initial appearance of the copper (II) chloride solution and of the Al strip.
3. If instructed to do so, thoroughly sand two 1 cm x 5 cm pieces of aluminum. Be careful, the edges are sharp.
4. Describe the appearance of both reactants in the Procedure 2 data table.
5. Once the water bath is boiling, remove the beaker from the heat source and place it on the ceramic-centered wire gauze on the bench. Turn off the hot plate and unplug it or turn off the burner.
6. Slide one aluminum strip into the copper (II) chloride solution test tube, the Rxn test tube. Push it down with a stirring rod if necessary. Slide the second aluminum strip into the water test tube, the Ref Al test tube.
7. Make initial observations and record your observations on the lab report. Place the Rxn Test tube and both Ref test tubes in the hot water bath. Make additional observations after the test tubes have been in the hot water bath for 1, 3 and 5 minutes.
8. After the 5 minute observation, take your glass rod and slowly slide it along both sides of the aluminum strip to knock off the copper that has formed.
9. Write the molecular, complete ionic, and net ionic equations for the reaction of aluminum with copper(II) chloride in the Analysis section.
10. Based on the complete ionic equation, use the element and ion cards to complete the test tube diagram on the Before and After Test Tube sheets. Place all cards in realistic locations (solid, gas, aqueous) on the Before and After Test Tube sheets.
11. Have the instructor check your Test Tube Diagram sheets.
12. After receiving instructor approval, complete the Test Tube Diagrams in the Analysis section and answer the questions for Procedure 2. Also answer any questions pertaining to Procedure 2 found in the Conclusions and Extensions section.
13. After 10 minutes, remove the test tube from the hot water bath and make final observations. Place the test tubes in your test tube rack to cool. After 5 minutes, or when the test tube is cool enough to be handled safely, slowly pour 5 mL of the solution from the “Rxn” test tube onto a watch glass labeled with your name. Place the watch glass at the location designated by your instructor. Discard the test tubes from this procedure into the appropriately labeled discard beaker. In the beaker labeled Discard Proc 2: Rxn, discard the remaining contents of the “Rxn” test tube. If the aluminum strip sticks to the walls, add water and quickly dump it into the discard beaker again. Pour the contents of the Ref CuCl₂ test tube in the beaker labeled Discard Proc 2: Ref CuCl₂ and pour the contents of the Ref Al test tube in its the beaker.

Procedure 1 Continued
1. If the magnesium strip did not dissolve earlier, it should be dissolved by now. Make final observations.
2. Quickly dump the contents of the test tube into the beaker labeled Discard Procedure 1. If necessary to remove any remaining magnesium strip, add water and quickly pour the contents into the discard beaker again.

Clean-up
1. Rinse both test tubes with water several times and clean them with soap and a brush.
2. Clean-up your lab area.
3. Complete the lab report.

Data
Provide detailed descriptions of the reactants and products as well as evidence of reaction in the tables below.

<table>
<thead>
<tr>
<th>Procedure 1: The Reaction between Magnesium (s) and Nitric Acid (aq)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial Appearance &amp; Temperature (Before Mixing)</td>
</tr>
<tr>
<td>Magnesium</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Procedure 1: The Reaction between Magnesium (s) and Nitric Acid (aq)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Appearance Immediately after Mixing</td>
</tr>
<tr>
<td>Nitric Acid &amp; Magnesium Test Tube (Reaction Test Tube)</td>
</tr>
<tr>
<td>Appearance at 2 Minutes</td>
</tr>
<tr>
<td>-------------------------</td>
</tr>
<tr>
<td>Temperature &amp; Appearance at 5 Minutes, Including How the Test Tube Felt to the Touch</td>
</tr>
<tr>
<td>Appearance at 10 Minutes</td>
</tr>
</tbody>
</table>

**Analysis (Procedure 1: Solid Magnesium and Nitric Acid Solution)**

1. Write the balanced molecular equation for the reaction between the magnesium strip and the nitric acid solution. Include physical states.

2. Write the balanced complete ionic equation for the reaction between the magnesium strip and nitric acid solution. Include physical states.

3. Write the balanced net ionic equation for the reaction between the magnesium strip and the nitric acid solution. Include physical states.

4. Complete the Before and After Test Tube Diagrams below using element and ion symbols.

Before Rxn

After Rxn
5. Let’s consider masses. Look at your Test Tube Diagram. Consider each particle as representing one mole of that species, two particles two moles, etc.

a. Calculate the total mass of all of the reactants in the Before Test Tube. Show the calculation.

b. Based on your answer to 5a, what should the total mass of products in the After Test Tube be? Why?

c. Calculate the total mass of all of the products in and above the After Test Tube. Show the calculation.

d. How do your answers to questions 5a & 5c illustrate the Law of Conservation of Mass?

Data
Provide detailed descriptions of the reactants and products as well as evidence of reaction in the tables below.

<table>
<thead>
<tr>
<th>Procedure 2: The Reaction between Aluminum (s) and Copper(II) Chloride (aq)</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
</tr>
<tr>
<td>Initial Appearance &amp; Temperature (Before Mixing)</td>
</tr>
<tr>
<td>Aluminum (s)</td>
</tr>
<tr>
<td></td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Procedure 2: The Reaction between Aluminum (s) and Copper(II) Chloride (aq)</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
</tr>
<tr>
<td>Appearance Immediately after Mixing</td>
</tr>
<tr>
<td>Aluminum &amp; Copper(II) Chloride Test Tube (Reaction Test Tube)</td>
</tr>
<tr>
<td></td>
</tr>
<tr>
<td>Appearance at 1 Minute</td>
</tr>
<tr>
<td></td>
</tr>
<tr>
<td>Appearance at 3 Minutes,</td>
</tr>
<tr>
<td></td>
</tr>
</tbody>
</table>
Analysis (Procedure 2: Solid Aluminum and Copper(II) Chloride Solution)

1. Write the balanced molecular equation for the reaction between the aluminum strip and the copper(II) chloride solution. Include physical states.

2. Write the balanced complete ionic equation for the reaction between the aluminum strip and the copper(II) chloride solution. Include physical states.

3. Write the balanced net ionic equation for the reaction between the aluminum strip and the copper(II) chloride solution. Include physical states.

4. Complete the Before and After Test Tube Diagrams below using element and ion symbols.

Before Rxn

After Rxn
5. Atoms are electrically neutral as the number of electrons is equal to the number of protons. Ions carry charges but are present in the compounds they comprise in numbers such that the total positive charge is equal to the total negative charge. Therefore, what net charge would you expect:

   a. In the Before Test Tube?
   b. In the After Test Tube?
   c. Throughout the reaction?

6. Let’s consider electrical charge in more detail.
   - Refer to your Test Tube equation Diagram for a-c.
   - Refer to your complete ionic equation for d-f.

   a. What is the total positive charge in the Before Test Tube?
   b. What is the total negative charge in the Before Test Tube?
   c. What is the net charge in the Before Test Tube?
   d. What is the total positive charge on the reactants side?
   e. What is the total negative charge on the reactants side?
   f. What is the net charge on the reactants side?

Conclusions and Extensions
1. Consider the first reaction between magnesium and nitric acid.

   a. Which species gave away or lost electrons?
   b. Which species was oxidized?
   c. How many electrons did each magnesium atom lose?
   d. Which species took in or gained electrons?
   e. Which species was reduced?
   f. How many electrons did each hydrogen ion gain?
   g. How many hydrogen ions were required to react with each magnesium atom?

2. A gas was produced in this reaction. What was the gas?

3. What color is elemental, metallic magnesium, Mg⁰?

4. What color is the magnesium ion, Mg²⁺, in solution?

5. What is the spectator ion in this reaction?

6. Consider the temperature and how the test tube felt after 5 minutes. Is the reaction between Mg and HNO₃ endothermic or exothermic?

7. Now consider the second reaction between aluminum and a solution of copper(II) chloride.

   a. Which species gave away or lost electrons?
   b. Which species was oxidized?
c. How many electrons did each aluminum atom lose?
d. How many electrons are lost by 2 aluminum atoms?
e. Which species took in or gained electrons?
f. Which species was reduced?
g. How many electrons did each copper(II) ion gain?
h. How many electrons are gained by 3 copper(II) ions?
i. What is the ratio of electrons lost to electrons gained? e\textsuperscript{-} lost : e\textsuperscript{-} gained
j. What is the aluminum to copper mole ratio in the net ionic equation? Al:Cu

8. What color is elemental, metallic aluminum, Al\textsuperscript{0}?
9. What color is the aluminum ion, Al\textsuperscript{3+}, in solution?
10. What color is the copper(II) ion, Cu\textsuperscript{2+}, in solution?
11. What color is elemental, metallic copper, Cu\textsuperscript{0}?
12. What is the spectator ion in this reaction?

13. The Kinetic Molecular Theory postulates that all particles - atoms, molecules and ions - are in constant motion. In the case of a solid, the particles are confined to vibrating about fixed positions. But in liquids, solutions, and gases the particles randomly move throughout the container. To represent this molecular motion, place your index fingers on the two aluminum atoms on the Before Test Tube sheet and move them apart slightly then back together, apart, together, apart, together, etc. All the atoms within a solid sample can do is vibrate about fixed positions.

Now place one index finger on a Cu\textsuperscript{2+} ion in the Before test tube. Ions in solution are free to randomly, independently, move throughout the solution. What happens when a Cu\textsuperscript{2+} ion happens, by virtue of its random motion, to collide with one of the Al atoms? And a second Cu\textsuperscript{2+} ion with a second Al atom? And a third Cu\textsuperscript{2+} ion with the Al atoms?

14. Consider Procedure 1, the reaction between magnesium and a solution of nitric acid.

a. What evidence was there that a chemical reaction occurred?

b. The reaction slowed down noticeably from the beginning to the middle to the end. Provide an explanation for this decrease in reaction rate. Hint: Picture H\textsuperscript{1+} ions in solution bumping into magnesium atoms on the strip. As reaction proceeds, what happens to the number of H\textsuperscript{1+} ions in solution?

c. Which reactant was completely consumed in this reaction? This reactant is called the limiting reactant as it places a limit on how much product, H\textsubscript{2} in this case, was produced.

d. If you poured some of the resulting colorless solution onto a watch glass and allowed the water to evaporate, what substance would you find on the watch glass after the water evaporates?
15. Consider the second reaction between solid aluminum and a solution of copper(II) chloride.

a. What evidence was there that a chemical reaction occurred?

b. What species is responsible for the blue color?

c. What happens to the concentration of Cu\(^{2+}\) (aq) (increases, decreases, or remains the same) as the reaction progresses?

d. What happens to the blue color as the concentration of Cu\(^{2+}\) (aq) decreases?

16. You poured some of the solution from Procedure 2, the reaction of aluminum with copper(II) chloride, onto a watch glass. After the water evaporates from the watch glass, what do you expect to remain on the watch glass? *Don’t forget to check your watch glass in a few days.

17. In this lab metals reacted with hydrogen or another metal. Metals are reactive. They have low ionization energies and readily give up valence electrons to another species often resulting in a noble gas electron configuration for the metal. Metals are easily oxidized. Metals form cations.

Oxygen is reactive. It has a large negative electron affinity and readily gains electrons achieving a noble gas electron configuration. Oxygen is easily reduced. Oxygen forms a 2- anion.

Metals love oxygen, oxygen loves metals. And this presents a common problem in the world known as oxidation of the metal or, in the case of iron, rusting.

Metals are common materials used in structures and objects in our world. Look up two ways that metals can be protected from oxidation.

a. State how each method prevents oxidation.

b. Identify one advantage and one disadvantage of each method.

c. Identify one common application of each method.

*Answer using complete sentences.
Extension: Writing Prompt
Single Displacement Reaction between Zinc and Copper(II) Sulfate

- \( Zn \ (s) = \text{silver/gray} \)
- \( Cu \ (s) = \text{red-brown} \)

The following rxn occurs:

\[
Zn \ (s) + \ CuSO_4 \ (aq) \rightarrow ZnSO_4 \ (aq) + Cu \ (s)
\]

1. Write the complete ionic equation for this reaction:

2. Complete the before & after beakers for this reactions. Which equation, molecular or complete ionic, best reflects reality? (Circle one)

3. Answer the following questions.
   a. You are a zinc atom on the surface of a zinc strip that has just been placed in a beaker containing a solution with the stoichiometric amount of copper(II) sulfate. Write a paragraph, using complete sentence, describing the changes that are occurring to you in your:
      - Location
      - Charge
      - Did you lose or gain electrons, or no change?
      - Were you oxidized or reduced?
      - Where did the electrons go to or come from, or no change?
      - Color
      - Physical state
      - Identify the particles surrounding you.

   **Zinc Atom response:**
b. Now you are a copper(I) ion floating around in the solution in the same beaker. Write a paragraph, using complete sentence, describing the changes that are occurring to you in your:

- Location
- Charge
- Did you lose or gain electrons, or no change?
- Were you oxidized or reduced?
- Where did the electrons go to or come from, or no change?
- Color
- Physical state
- Identify the particles surrounding you.

*Copper (II) Ion response:*

c. Now you are a sulfate ion in the solution in the same beaker. Write a paragraph, using complete sentence, describing the changes that are occurring to you in your:

- Location
- Charge
- Did you lose or gain electrons, or no change?
- Were you oxidized or reduced?
- Where did the electrons go to or come from, or no change?
- Color
- Physical state
- Identify the particles surrounding you.

*Sulfate Ion response:*