

Metal Activity and Reactivity

Oxidation and Reduction Reactions

Introduction

The usefulness of metals in structural and other applications depends on their physical and chemical properties. Although iron is the most common metal used in manufacturing, it must be protected against corrosion because iron rusts easily. Copper is used in electrical wiring because it conducts electricity extremely well and resists corrosion better than many metals. Gold is a highly valuable jewelry metal because it is essentially unreactive. How can we determine the relative reactivity of different metals?

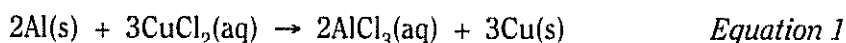
Concepts

- Activity series
- Single replacement reactions
- Oxidation–reduction
- Half-reactions

Background

A ranking of metals in order of their tendency to react with acids and water is called an *activity series*. Comparing the reactions of sodium, magnesium, and aluminum reveals that sodium reacts violently with acids and water, magnesium reacts with acids and hot water, and aluminum reacts only with acids. Based on the trend in reactivity, sodium is more active than magnesium, which is more active than aluminum. The activity series for these metals would be written as $\text{Na} > \text{Mg} > \text{Al}$.

Another way to determine the activity of metals is to compare the reactions of metals with different metal ions. Consider, for example, what happens when a piece of aluminum foil is placed in a solution of copper(II) chloride. A vigorous reaction is observed—heat is released, the blue color due to copper(II) ions fades, the aluminum foil disintegrates, and a new, red-dish brown solid appears in the reaction mixture. The reaction is summarized in Equation 1. Copper(II) ions from copper(II) chloride are converted to copper metal, and aluminum metal is converted to aluminum cations in aluminum chloride, which is soluble in water. In contrast, when a piece of copper metal is placed in a solution of aluminum chloride, no reaction takes place (Equation 2).



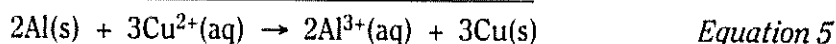
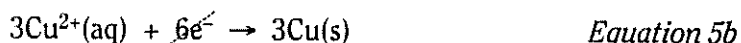
The reaction of aluminum with copper(II) chloride is classified as a *single replacement* reaction—aluminum reacts with and “replaces” copper ions in copper(II) chloride. Single replacement reactions will occur spontaneously in one direction only (compare Equations 1 and 2). *A more active metal always replaces the ion of a less active metal.* In general, the activity of a metal may be defined as follows: An active metal will react with a compound of a less active metal, which is converted to its “free element” form. The more active metal forms a new compound containing metal cations. Based on Equation 1, aluminum is more active than copper.

Single replacement reactions are examples of *oxidation–reduction reactions*. *Oxidation* is defined as the process of losing electrons, and a substance that loses electrons during a chemical reaction is said to be oxidized. If one substance loses electrons during the course of a chemical reaction, another substance must gain electrons. The process of gaining electrons is called *reduction*, and a substance that gains electrons during a chemical reaction is said to be reduced. Oxidation and reduction occur together so that the total number of electrons lost by the substance that is oxidized will be equal to the number of electrons gained by the substance that is reduced. The number of moles of each reactant in the balanced chemical equation for an oxidation–reduction reaction reflects “electron balance” as well as “atom balance.”

The loss and gain of electrons by the reactants in Equation 1 will be more apparent if the overall reaction is broken down into two separate half-reactions. Equations 3 and 4 show the oxidation and reduction half-reactions, respectively, for the reaction of aluminum metal with copper(II) ions. In the oxidation half-reaction (Equation 3), each aluminum atom loses three electrons and is oxidized to an Al^{3+} ion. In the reduction half-reaction (Equation 4), each Cu^{2+} ion gains two electrons and is reduced to a copper atom. Notice that the total charge on both sides of a half-reaction must be the same (charge is conserved). The chloride ions remain dissolved in solution and are not shown in the half-reactions.



The number of electrons involved in the overall oxidation–reduction reaction will be balanced if Equation 3 is multiplied by a factor of two (Equation 5a) and Equation 4 is multiplied by a factor of three (Equation 5b). When the resulting half-reactions are added together, the number of electrons lost by the aluminum atoms is equal to the number of electrons gained by the copper ions and the electrons “cancel out” of the overall equation. Equation 5 is the balanced, *net ionic equation* for the reaction of aluminum with copper(II) chloride—the chloride ions are “spectator ions” and are not shown.



Experiment Overview

The purpose of this experiment is to carry out a series of possible single replacement reactions of metals with solutions of metal cations and to determine the activity series of the metals. The more active metal in each pair will be determined by observing which combinations of metals and metal cations undergo a chemical reaction. The metals will be ranked from most active to least active.

Pre-Lab Questions

1. According to the *Background* information, magnesium is more active than aluminum. Predict which combination of metal and metal ion solution will show evidence of a chemical reaction: (a) Mg(s) and $\text{Al}_2(\text{SO}_4)_3(\text{aq})$ or (b) Al(s) and $\text{MgSO}_4(\text{aq})$.
2. Write the balanced chemical equation for the single replacement reaction that is expected to occur (1a or 1b).
3. Which substance is being oxidized in this reaction? Which substance is being reduced?
4. Write the oxidation and reduction half-reactions for the overall reaction.

Materials

Copper, Cu, 1-cm ² strips, 5	Distilled water and wash bottle
Copper(II) sulfate solution, CuSO_4 , 0.2 M, 4 mL	Forceps or tweezers
Hydrochloric acid, HCl, 2 M, 4 mL	Magnifier (optional)
Iron, Fe, 1-cm ² strips, 5	Marking pen
Iron(II) sulfate solution, FeSO_4 , 0.2 M, 4 mL	Paper towels
Magnesium ribbon, Mg, 1-cm, 5	Pipets, Beral-type, 6
Magnesium nitrate solution, $\text{Mg}(\text{NO}_3)_2$, 0.2 M, 4 mL	Reaction plate, 24-well
Silver nitrate solution, AgNO_3 , 0.2 M, 4 mL	Sandpaper (optional)
Zinc foil, Zn, 1-cm ² squares, 5	Scissors
Zinc sulfate solution, ZnSO_4 , 0.2 M, 4 mL	Toothpicks (optional)

Safety Precautions

Hydrochloric acid is a corrosive liquid and toxic by ingestion or inhalation. Silver nitrate is slightly toxic by ingestion and will stain skin and clothing. Copper(II) nitrate and iron(II) sulfate are toxic by ingestion. Metal pieces may have sharp edges—handle with care. Avoid contact of all chemicals with eyes and skin. Wear chemical splash goggles, chemical-resistant gloves, and a chemical-resistant apron. Wash hands thoroughly with soap and water before leaving the lab.

Procedure

1. Place a 24-well reaction plate on top of a sheet of white paper as shown below. Each well is identified by a unique combination of a letter and a number, where the letter refers to a horizontal row and the number to a vertical column (Figure 1).

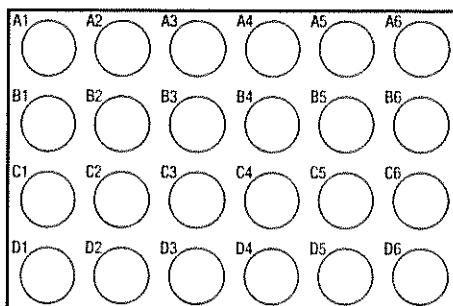


Figure 1. Layout and Numbering of a 24-Well Reaction Plate.

2. Obtain or cut five 1-cm² pieces of each metal to be tested—copper, iron, magnesium and zinc—and lay out the metals in labeled rows on a paper towel. Follow your instructor's directions for polishing the metals, if necessary.

Consult the data table frequently and carefully read each label. Notice that it is not necessary to test a metal with a solution of its own cation, e.g., copper is not tested in copper(II) sulfate, etc.

3. Using a clean, labeled pipet or eyedropper for each solution, add 15 drops of the appropriate solution to each well, as follows:
 - Copper(II) sulfate solution to each well A1–D1
 - Iron(II) sulfate solution to each well A2–D2
 - Magnesium nitrate solution to each well A3–D3
 - Silver nitrate solution to each well A4–D4
 - Zinc sulfate solution to each well A5–D5
 - Hydrochloric acid solution to each well A6–D6
4. Place one piece of copper into each well A2–A6.
5. Place one piece of iron into well B1 and into each well B3–B6.
6. Place one piece of magnesium into wells C1 and C2 and into each well C4–C6.
7. Place one piece of zinc into each well D1–D4 and into well D6.
8. After all the metal pieces have been added, observe and record any evidence of a chemical reaction in each well. Use a clean toothpick to submerge any metal pieces if necessary.
9. Some chemical reactions may be spontaneous but slow—continue to observe the reactions for 3–5 minutes.
10. Sometimes the only sign of a chemical reaction is a slight darkening or coating of the metal surface. Use clean forceps and a magnifier to observe any metals where the evidence of reaction is not obvious. It may be helpful to compare the color and appearance of a test metal with a clean, untreated piece of metal. Record all observations in the data table.
11. If no evidence of a reaction is observed in a well, write NR (no reaction). Remember, single replacement reactions will occur in one direction only. Many of the metal/metal ion combinations will show no reaction!

Disposal

12. Do NOT dispose of metal pieces down the drain. Fold several paper towels together to form a thick mat. Gently shake the contents of the reaction plate onto the folded paper towels. Use forceps to remove any metal pieces sticking to the bottom or sides of the reaction plate. The folded paper towels may be discarded in the trash.
13. Use cotton swabs to clean out any remaining residues in the reaction plate. Wash the reaction plate thoroughly under running water and rinse well with distilled water.

Name: _____

Class/Lab Period: _____

Metal Activity and Reactivity

Data Table

	CuSO_4	FeSO_4	$\text{Mg}(\text{NO}_3)_2$	AgNO_3	ZnSO_4	HCl
Cu						
Fe						
Mg						
Zn						

Post-Lab Questions *(Answer in complete sentences where appropriate.)*

1. Which metal reacted with the most metal ion solutions?
2. Which metal reacted with the fewest metal ion solutions?
3. Compare the general trend in the reactivity of Cu, Fe, Mg, and Zn and rank the metals from most active (first) to least active (last).
4. Because silver metal is expensive, it was not used in this experiment. Based on the observed reactions of Cu, Fe, Mg, and Zn with silver nitrate, explain why it was not necessary to test silver in order to determine its activity.

5. Rewrite the activity series of the metals (Question #3) to include silver.
6. What evidence was obtained to show that some of the metals reacted with hydrochloric acid?
7. Acids are sources of hydrogen ions, H^+ or H_3O^+ . A metal is considered more active than hydrogen if it will react with acids. Rewrite the activity series of the metals (Question #5) to include hydrogen.
8. Write the balanced chemical equation for the reaction of the *most active metal* tested with hydrochloric acid. Identify the substance that was oxidized and the substance that was reduced in this oxidation–reduction reaction. *Hint:* Write the net ionic equation for the reaction.
9. Write the balanced, *net ionic equation* for each observed reaction of a metal with a metal ion. *Hint:* Substitute the symbols of the metals and metal ions into the following equation and remember to balance the charges.
$$\text{M(s)} + \text{N}^{x+}(\text{aq}) \rightarrow \text{M}^{y+}(\text{aq}) + \text{N(s)}$$
10. (*Optional*) The substance that accepts electrons and is reduced in an oxidation–reduction reaction is called the oxidizing agent (it causes the oxidation of another substance). The substance that gives up electrons and is oxidized acts as a reducing agent (it causes the reduction of another substance). An active metal is a good reducing agent: *True or False?* Explain.