

Activity Series Performer's Version

Safety Hazards

- Personal Protective Equipment:
 - Safety glasses/goggles
 - Nitrile gloves
 - Chemical & flame retardant lab coat
- Physical Hazards
 - Anhydrous copper(II) sulfate is very toxic to aquatic life with long lasting effects.
 - Silver nitrate may be corrosive to metals and is an oxidizer that may intensify fire. Silver nitrate is very toxic to aquatic life with long lasting effects.
 - Aluminum sulfate x-hydrate is corrosive to metals.
- Chemical Hazards
 - Anhydrous copper(II) sulfate is toxic if swallowed and causes skin irritation and serious eye irritation.
 - Silver nitrate causes severe skin burns and eye damage and may damage fertility or the unborn child.

- Aluminum sulfate x-hydrate causes serious eye damage.

Materials

- 7.98g Copper sulfate, anhydrous
- 8.58g Silver nitrate, 99%
- 33.32g Aluminum sulfate, x-hydrate
- 15g Sodium chloride (optional)
- 1.5" – 2" Silver wire
- 2x 1.5" – 2" Copper wire
- 1.5" in² Aluminum foil
- 4x small petri dishes
- Polypad

Safety Data Sheet(s)

- [Copper sulfate, anhydrous](#)
- [Silver nitrate, 99%](#)
- [Aluminum sulfate, x-hydrate](#)
- [Sodium chloride](#)
- [Silver wire](#)
- [Copper wire](#)
- [Aluminum foil](#)

Procedure

1. Prepare the petri dishes (but don't place the wire and foil into the solutions yet) and equally space them :
 - a. Dish #1: Silver wire in 0.5M Copper(II) sulfate solution
 - b. Dish #2: Copper wire in 0.5M Silver nitrate solution
 - c. Dish #3: Copper wire in 0.5M Aluminum sulfate solution
 - d. Dish #4: Aluminum foil in 0.5M Copper(II) sulfate solution
2. When you're ready, place each wire/foil piece into the solutions and observe.
 - a. Note: If you want to kick the aluminum foil reaction into gear, sprinkle some sodium chloride into the petri dish!

Pedagogy & Supplemental Information

In this demonstration, four petri dishes are placed under a document camera to showcase the reactivity of metals through single replacement reactions. Each dish contains a strip of metal submerged in a 0.5M solution of a different metal ion. The goal is to observe which combinations produce visible redox reactions and to interpret these results using the activity series – a ranking of metals based on how readily they lose electrons to form positive ions.

The activity series reflects each metal's tendency to be oxidized and is rooted in standard reduction potentials, which quantify how easily a species gains electrons under standard conditions. A metal with a lower (more negative) reduction potential is more likely to give up electrons and be oxidized, making it more reactive. In contrast, a metal ion with a higher (more positive) reduction potential is more likely to accept electrons and be reduced. A redox reaction is favorable when a more reactive metal (higher in the activity series) donates electrons to a less reactive metal ion.

This principle is clearly illustrated in the reaction between copper wire and silver nitrate. As the copper is oxidized to Cu^{2+} , silver ions are reduced to elemental silver, which visibly deposits on the copper wire as a gray, crystalline layer. This occurs because copper is above silver in the activity series, meaning it is more easily oxidized. In contrast, silver wire in copper(II) sulfate shows no reaction; silver is less reactive than copper and cannot reduce Cu^{2+} ions, so the solution and the wire remain unchanged.

A similar lack of reactivity is observed with copper wire in aluminum sulfate. Copper, again, is lower in the activity series than aluminum and therefore cannot displace aluminum from solution. However, placing aluminum foil in copper(II) sulfate results in a more complex behavior. Initially, little reaction is observed due to aluminum's natural oxide coating, which passivates the surface and prevents electron transfer. When a small amount of sodium chloride is added, chloride ions disrupt the oxide layer, exposing bare aluminum. Once exposed, aluminum – among the most reactive common metals – quickly reduces Cu^{2+} ions to elemental copper metal, which deposits on the foil while gas bubbles evolve at the surface. This reaction is notably slower and less vigorous than the similar "Disappearing Aluminum Rod" demonstration with copper(II) chloride. The difference arises from chloride's dual role: it helps break the oxide layer and forms soluble complexes with aluminum ions, enhancing dissolution and overall reactivity. Sulfate ions, in contrast, do not support this behavior as effectively.

This demonstration reinforces the key idea that redox reactivity is not absolute, but rather is always relative to the species involved. Whether or not a reaction occurs depends on the ability of the solid metal to act as a reducing agent for the metal ions in solution. Both the activity series and standard reduction potentials offer predictive frameworks for these outcomes.

The concepts illustrated here have wide-reaching applications. In corrosion science, understanding how metals displace one another helps in selecting materials and designing protective coatings. In battery technology, redox pairs are selected based on their positions in the activity series to drive electron flow. Metal refining, electroplating, and even water treatment systems rely on these same electrochemical principles. Through this simple but informative demo, students connect visual changes with fundamental redox chemistry that powers much of the modern world.