Chemistry of Copper

Lab Session 2

Pages 59 – 66

Pre-lab (all questions) is page 63

All lab questions, page 66 – due by the start of your next lab session

Introduction

- Copper is found in group 1, MW = 63.456
- Shiny (orange/red color), Malleable, Ductile
- Oxidizes in air (turns a green color – patina)
- Oxidation States are Cu(I) = $3d^{10}$, Cu(II) =$3d^{9}$
- Cu(II) compounds are usually blue or green in color

Lab Objectives

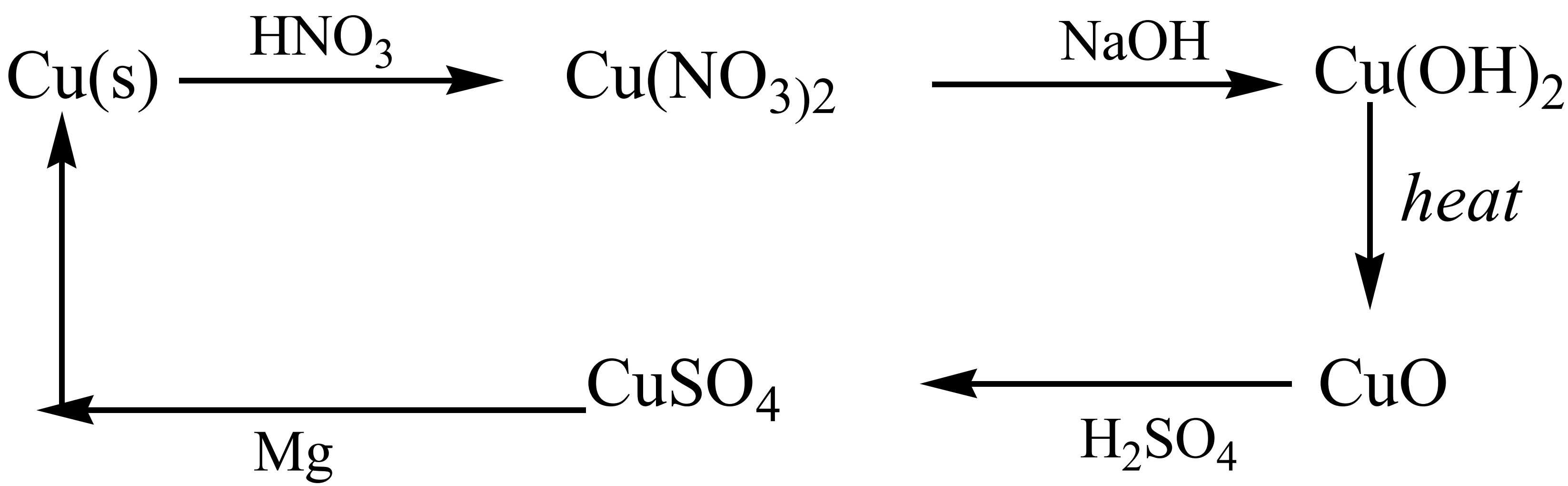
- Observation of Copper's Chemical Properties
- Isolation of several copper compounds
- Determination of percent recovery of Cu

- Work in partners
- Complete pre-lab prior to lab session
- Lab time is expected to be 3 hours

- In this experiment you will take a copper sample through a series of five reactions

- The end product will be your original copper sample, making this a cycle of reactions.

- With careful attention to quantitative lab practices, you should be able to recover all the copper you started with.


Cu(s)
HNO3
Cu(NO3)2
NaOH
Cu(OH)2
heat
CuO
H2SO4
CuSO4
Mg

What will you do?

- Perform each reaction and write your observations.
  - Color change, precipitate, gas evolved
- Use equations to interpret observations
- Retain as much copper as possible

Think about your reactions

- What kind of reaction is it?
  - Acid/base, gas evolved, redox, precipitate
- What is your reactant?
- How do you know if the reaction is complete?
- How can you minimize loss of Cu and its compounds?

Types of Reactions

- Combustion
- Synthesis
- Single Displacement
- Double Displacement (Metathesis)
- Decomposition
- Acid – Base

(single, double displacement and acid base, can also be redox or double displacement and acid base)

Combustion

- A combustion reaction is when oxygen combines with another compound to form water and carbon dioxide. These reactions are exothermic, meaning they produce heat.
- An example of this kind of reaction is the burning of naphthalene:

$$\mathbf{C_{10}H_8 + 12\ O_2 \rightarrow 10\ CO_2 + 4\ H_2O}$$

Decomposition Reaction

- A more complex substance breaks down into its more simple parts. One reactant yields 2 or more products.
- For example, water can be broken down into hydrogen gas and oxygen gas.
- $H_2O \rightarrow \frac{1}{2}\, O_2 + H_2$
- The chemical equation for a decomposition reaction looks like:

$$\text{reactant} \rightarrow \text{product} + \text{product}$$

Synthesis

- A synthesis reaction is when two or more simple compounds combine to form a more complicated one. These reactions come in the general form of:
  - A + B ---> AB
- One example of a synthesis reaction is the combination of iron and sulfur to form iron (II) sulfide:
  - 8 Fe + S_8 ---> 8 FeS

Single displacement

- This is when one element trades places with another element in a compound. These reactions come in the general form of:

$$A + BC \rightarrow AC + B$$

- One example of a single displacement reaction is when magnesium replaces hydrogen in water to make magnesium hydroxide and hydrogen gas:

$$Mg\ (s) + 2\ H_2O\ (l) \rightarrow Mg(OH)_2\ (s) + H_2\ (g)$$

Double displacement

- This is when the anions and cations of two different molecules switch places, forming two entirely different compounds. These reactions are in the general form:

$$AB + CD \rightarrow AD + CB$$

- One example of a double displacement reaction is the reaction of lead (II) nitrate with potassium iodide to form lead (II) iodide and potassium nitrate:

$$Pb(NO_3)_2 + 2\ KI \rightarrow PbI_2 + 2\ KNO_3$$

Acid-base

- This is a special kind of double displacement reaction that takes place when an acid and base react with each other. The H^+ ion in the acid reacts with the OH^- ion in the base, causing the formation of water. Generally, the product of this reaction is some ionic salt and water:

$$\mathbf{HA + BOH \rightarrow H_2O + BA}$$

- One example of an acid-base reaction is the reaction of hydrobromic acid (HBr) with sodium hydroxide:

$$\mathbf{HBr + NaOH \rightarrow NaBr + H_2O}$$

What type of reaction is taking place?

Follow this series of questions. When you can answer " yes: " to a question, then stop.

1) Does your reaction have oxygen as one of it's reactants and carbon dioxide and water as products?

   Yes: it's a combustion reaction

2) Does your reaction have two (or more) chemicals combining to form one chemical?

   Yes: it's a synthesis reaction

3) Does your reaction have one large molecule falling apart to make several small ones?

   Yes: it's a decomposition reaction

4) Does your reaction have any molecules that contain only one element?

Yes: it's a single displacement reaction

5) Does your reaction have water as one of the products?

Yes : it's an acid-base reaction

6) If you haven't answered " yes " to any of the questions above, then you've got a double displacement reaction

- (1) $8HNO_{3\,(aq)} + 3Cu_{(s)} + O_{2\,(g)} \rightarrow 3Cu(NO_3)_{2\,(aq)} + 4H_2O_{(l)} + 2NO_{2\,(g)}$

- (2) $Cu(NO_3)_{2\,(aq)} + 2NaOH_{(aq)} \rightarrow Cu(OH)_{2\,(s)} + 2NaNO_{3\,(aq)}$

- (3) $Cu(OH)_{2\,(s)}\ (heat) \rightarrow CuO_{(s)} + H_2O_{(l)}$

- (4) $CuO_{(s)} + H_2SO_{4\,(aq)} \rightarrow CuSO_{4\,(aq)} + H_2O_{(l)}$

- (5) $CuSO_{4\,(aq)} + Mg_{(s)} \rightarrow MgSO_{4\,(aq)} + Cu_{(s)}$

Copper Nitrate

$$8HNO_{3\ (aq)} + 3Cu_{\ (s)} + O_{2\ (g)} \rightarrow 3Cu(NO_3)_{2\ (aq)} + 4H_2O_{\ (l)} + 2NO_{2\ (g)}$$

- Crystalline $Cu(NO_3)_2(H_2O)_{2.5}$ features octahedral Cu centers surrounded by water and the nitrate anions.
- This hydrate decomposes at *ca*. 170 °C into copper(II) oxide, nitrogen dioxide and oxygen:
  - $2Cu(NO_3)_{2(s)} \rightarrow 2CuO_{(s)} + 4NO_{2(g)} + O_{2(g)}$

Precautions

- Conc. HNO_3 is corrosive
- Use a fume-hood – $NO_2(g)$ is generated which is toxic if inhaled
- Observe the color of the gas evolved
- Only had enough nitric acid so as your copper dissolves (more is not better!)

Use of Copper Nitrate

- Copper nitrate can be used to generate nitric acid by heating it until decomposition and passing the fumes directly into water. This method is similar to the last step in the Ostwald process. The equations are as follows:
  - $2Cu(NO_3)_2 \rightarrow 2CuO + 4NO_2\ (g) + O_2\ (g)$
  - $NO_2 + H_2O \rightarrow 2\ HNO_3 + NO\ (g)$
- Copper nitrate soaked splints of wood burn with an emerald green flame. Addition of Magnesium nitrate gives a lime green color.

Copper Nitrate to Copper Hydroxide

$$Cu(NO_3)_{2\ (aq)} + 2NaOH_{\ (aq)} \rightarrow Cu(OH)_{2\ (s)} + 2NaNO_{3\ (aq)}$$

The driving force of the reaction is the formation of a precipitate.

Sodium Hydroxide is a strong base – use with caution.

- The precipitate is separated through using a centrifuge.

- Ensure complete precipitation has occurred by adding a few extra drops of NaOH after using the centrifuge.

- The liquid above the precipitate is called the supernatant

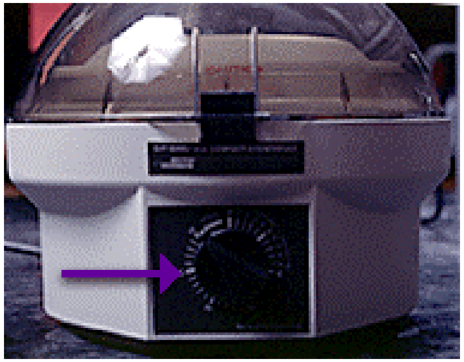
Lab Techniques

- Using a centrifuge

- A centrifuge separates a heterogeneous mixture of solid and liquid by spinning it. After a successful centrifugation, the solid precipitate settles to the bottom of the test tube and the solution, called the supernatant (centrifugate), is clear.

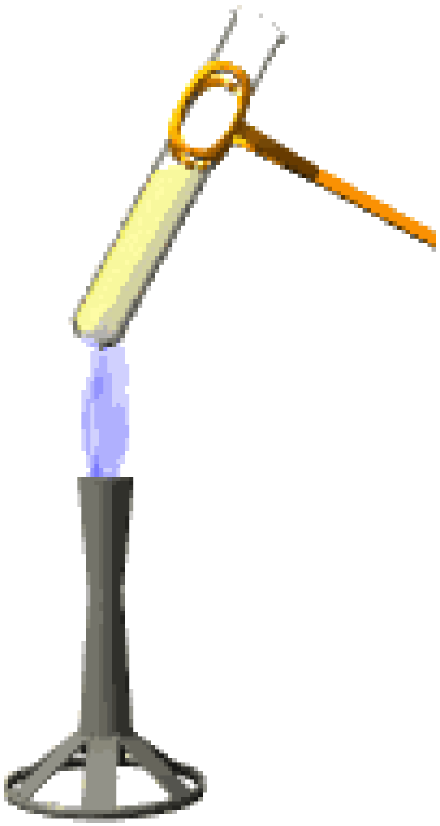
Place test tube in centrifuge holder.

Balance with another test tube filled to the same level in the opposite holder.



Close cover and turn on. Centrifugation takes a minute or more. Note that you must turn off the centrifuge with the switch and wait for it to stop spinning, to effectively separate the precipitate and solution.

Copper Hydroxide to Copper Oxide



$$Cu(OH)_{2\,(s)}\ (\textit{heat}\ ) \rightarrow CuO_{\,(s)} + H_2O_{\,(l)}$$

- Never heat a closed container, and be sure that open test tubes point away from you and others while being heated. Always heat the test tube at an angle from the flame.

Copper Oxide

- Copper(II) oxide is a basic oxide, so it dissolves in mineral acids such as hydrochloric acid, Sulfuric acid or nitric acid to give the corresponding copper(II) salts:

$$CuO + 2\ HNO_3 \rightarrow Cu(NO_3)_2 + H_2O$$

- It can also be reduced to copper metal using hydrogen or carbon monoxide:

$$H_2 + CuO \rightarrow Cu + H_2O$$

- Copper (II) oxide has uses as semiconductor

Copper Oxide to Copper Sulfate

$$CuO_{(s)} + H_2SO_{4\,(aq)} \rightarrow CuSO_{4\,(aq)} + H_2O_{(l)}$$

- Add approx 1ml of H_2SO_4 to the copper oxide
- The solution should change color.
- Record your observations.
- Add enough acid until all your oxide is dissolved.
- Slight heating may be required

Copper Sulfate

- This has the formula **$CuSO_4$**, and is a common salt of copper.
- Copper sulfate exists as a series of compounds that differ in their degree of hydration.
- The anhydrous form is a pale green or gray-white powder, while the hydrated form is bright blue. The archaic name for copper(II) sulfate is **blue vitriol**

Synthesis

- It is made by the action of sulfuric acid on a variety of copper(II) compounds, such copper(II) oxide and copper carbonate.
- Such reactions are considered acid-base reactions.
- Copper sulfate most occurs in nature as the pentahydrate ($CuSO_4 \cdot 5H_2O$).
- This mineral is called chalcanthite.

Uses

- Copper sulphate is also used to test blood for anemia. A drop of the patient's blood is dropped into a container of copper sulfate, if it sinks within a certain time, then the patient has sufficient hemoglobin levels and is not anemic. If the blood floats or sinks too slowly, then the patient is iron-deficient and may be anemic.
- In a flame test, copper ions emit a deep blue-green light, much more blue than the flame test for barium.

More Uses

- Copper(II) sulfate is a desiccant.
- Copper sulfate is a commonly included chemical in children's chemistry sets and is often used in high school crystal growing and copper plating experiments.
- A very dilute solution of Copper sulfate is used to treat aquarium fish of various parasitic infections. However, as the copper ions are also highly toxic to the fish, care must be taken with the dosage.

Copper Sulfate to Copper

- $CuSO_{4\,(aq)} + Mg_{(s)} \rightarrow MgSO_{4\,(aq)} + Cu_{(s)}$
- Why Magnesium metal?

- Magnesium is a more reactive metal than copper
- It is an alkali earth metal that is found in group 2
- It is a reducing agent and is therefore oxidized itself
- $Mg(s) \rightarrow Mg^{2+} + 2e'$

- Whenever something is oxidized, something else in the reaction needs to be reduced.
- The Cu^{2+} (as $CuSO_4$) picks up the 2 electrons forming Cu(s) and $MgSO_4$
- This is a single displacement reaction

$$A + BC \rightarrow AC + B$$

- Cu reacts readily with oxidizing agents.
- Oxidizing agents are reduced themselves

OILRIG

Oxidation Is Loss Reduction Is Gain

In chemical reactions, whenever an oxidation occurs a reduction is also present (and vice versa).

Recap: (page 218 Brady)

- The oxidation no. of a free element is 0.
- Ox. No. for a simple monoatomic ion, eg. Na^+, Cl^- is = to charge on ion.
- The sum of all the oxidation numbers of the atoms in a molecule or polyatomic ion must equal the charge on the particle.
- Fluorine has an ox. state of –1.
- H has an ox. state of +1 except when with electropositive elements then (-1)
- Oxygen usually has an oxidation state of –2.

$$Cu(s) + 4HNO_3(aq) \longrightarrow Cu(NO_3)_2\,(aq) + 2NO_2(g) + 2H_2O(l)$$

$$Cu(s) \longrightarrow Cu^{2+} + 2e^-$$ Copper is oxidized

To determine which species is reduced, the oxidation states of the other reactants and products are considered.

HNO_3 Oxidation States: H = +1, O = -2 (= -6) N = +5
NO_2 Oxidation States: N = +4 , O = -2

$$2HNO_3(aq) \longrightarrow 2NO_2$$

There must be an equal no. of electrons 'exchanged'.
The equation is balanced by adding H_2O

- An example of a redox reaction occurs between hydrochloric acid and zinc metal, where the Zn atoms lose electrons and are oxidized to form Zn^{2+} ions:
- $Zn(s) \text{ --> } Zn^{2+}(aq) + 2e^-$
- The H^+ ions of the HCl gain electrons and are reduced to H atoms, which combine to form H_2 molecules:
- $2H^+(aq) + 2e^- \text{ --> } H_2(g)$
- The overall equation for the reaction becomes:
- $Zn(s) + 2H^+(aq) \text{ --> } Zn^{2+}(aq) + H_2(g)$

Two important principles apply when writing balanced equations for reactions between species in a solution:

1. The balanced equation only includes the species that participate in forming products.

- For example, in the reaction between $AgNO_3$ and NaCl, the NO_3^- and Na^+ ions were not involved in the precipitation reaction and were not included in the balanced equation.

2. The total charge must be the same on both sides of a balanced equation.

- Note that the total charge can be zero or non-zero, as long as it is the same on both the reactants and products sides of the equation.

% of Copper Recovered

- The amount of Cu that you start with is recorded = Mi = Initial Mass
- The amount recovered is recorded, Mf = Final Mass
- % Recovered = Mi/Mf x 100