

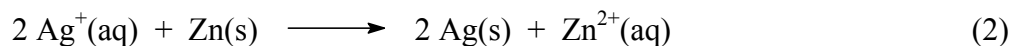
Oxidation-Reduction Reactions

Introduction

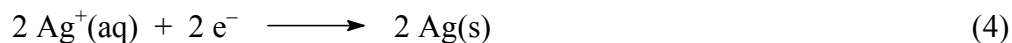
Many chemical reactions proceed by the electron (e^-) transfer from one reactant to another. These electron-transfer reactions are referred to as oxidation-reduction or redox reactions. Oxidation is defined as the part of a redox reaction in which a species loses electrons and increases in oxidation number. Reduction is the part of the redox reaction in which a species gains electrons and decreases in oxidation number. The oxidation number is related to the charge an atom has in a substance. In any reaction in which oxidation occurs, reduction must also occur. An oxidizing agent is a species that oxidizes another species and is itself reduced. A reducing agent is a species that reduces another species and is itself oxidized. We will use the example of zinc metal being placed in a silver nitrate solution. The molecular equation is



The net ionic equation is



You must remember that the oxidation number of an atom in an elementary substance is zero. It can be seen in equation 2 that the zinc metal loses two electrons to form the zinc(II) ions, while the two silver ions each gain one electron to form two silver metal atoms. Since this reaction involves the transfer of electrons it is a redox reaction. Equation 2 can be written in terms of two half-reactions. A half-reaction is one of two parts of a redox reaction, one of which involves a loss of electrons and the other which involves the gain of electrons. The half-reactions for equation 2 are



The number of electrons released in any oxidation half-reaction must equal the number of electrons consumed in the corresponding reduction half-reaction. Since the number of electrons is exact, the electrons do not appear in the overall redox reaction equation (2).

Procedure

1. Reaction of Cu^{2+} with Fe

- Add about 2 mL of 0.5 M CuSO_4 solution to a test tube.
- Add a small piece of steel wool (Fe) to the CuSO_4 solution and shake for a few minutes.
- Decant the **colorless** solution into a clean test tube to be used in Reaction 2.

2. Reaction of Fe^{2+} with H_2O_2

- Add 4 drops of 3% hydrogen peroxide to the colorless solution above.

To help characterize the product, acidify the solution with 5 drops of 6 M HNO_3 , then add 1 drop of 0.1 M KSCN. Consult a table of standard reduction potentials for the redox half-reaction of H_2O_2 .

3. Reaction of Cu^{2+} with Zn

- Add about 2 mL of 0.5 M CuSO_4 solution to a test tube.
- Add a small piece of zinc metal to the CuSO_4 solution and shake for a few minutes.
- After a fair amount of reddish-brown solid has formed and flaked off the piece of zinc metal, pour all of the solution and solids into a small beaker; rinsing if necessary.
- Remove the piece of zinc with tweezers and set it aside for use in Reaction 6. Transfer the reddish-brown solid into a clean crucible and remove as much water as possible by decanting. Use the crucible and solid in reaction 5.

Light your Bunsen burner and adjust the airflow to achieve a hot flame. Use it in the following **two** reactions.

4. Reaction of Mg with O_2

- Obtain a piece of Mg ribbon. Grasp the Mg ribbon with **crucible tongs** and hold the Mg in the flame until it ignites. Do not look **directly** at the reaction until it is over.

5. Reaction of Cu with O_2

- Place the crucible from Reaction 3 on a clay triangle supported by a ring and ring stand. Adjust the ring height to allow the hottest part of the Bunsen burner flame to just touch the bottom of the crucible.
- Gently** heat the contents of the crucible until **dry**, then strongly for five minutes. Carefully stir the contents of the crucible with a stirring rod and heat again for two minutes. Allow to cool.

To help characterize the **product**, transfer the solid from the crucible to a test tube and add ten drops of 10% HCl. add 10 drops of concentrated ammonia to the solution.

6. Reaction of Zn with H⁺

- a. Place the zinc saved from Reaction 3 in a test tube and add about 2 mL of 10% HCl. Observe the reaction, then dilute with distilled water to slow the reaction and discard.

7. Reaction of Mg with H⁺

- a. Obtain a piece of Mg ribbon and place in a test tube. Add about 2 mL of 10% HCl.

8. Reaction of Cl₂ with I⁻

- a. Place about 1 mL of 0.1 M KI solution in a test tube.
b. Add aqueous Cl₂ to the test tube until a distinct yellow solution is obtained.

In order to characterize the product of the reaction above, add a drop or two of 1 M Na₂S₂O₃ solution. Note what happens **initially**.

Question

The voltage of an alkaline battery is supplied by the following (unbalanced) reaction:



Identify the substance **oxidized** and the substance **reduced**.

Data Treatment and Discussion

1. Identify the product of all of the experiments performed in the laboratory and write the (net ionic) half-reactions. Identify each half-reaction as an oxidation or a reduction.
2. Combine the half-reactions of each experiment to give the overall, net ionic equations.

Conclusion

Address the following **based** on the experiments you have performed in this laboratory and **your observations**:

1. Which is a stronger reducing agent, Cu or Zn? Explain.
2. Which is a stronger oxidizing agent, Cl₂ or I₂? Explain.
3. Which is a stronger reducing agent, Fe or Cu? Explain.
4. A bottle of FeSO₄ **solution** was accidentally left open to the **air**. The solution has turned a pale yellow color. What happened (chemically)?