Oxidation–Reduction Survey
Iron(II) and Iron(III) Reactions

Introduction

Iron exists in the body in two forms—iron(II), Fe^{2+}, and iron(III), Fe^{3+} ions. Interconversion of the two forms of iron takes place via the loss or gain of an electron. Let's investigate the role of electron transfer in the oxidation and reduction of iron(II) and iron(III) compounds, respectively.

Concepts

- Oxidation–reduction
- Oxidation state
- Half-reactions
- Oxidizing and reducing agents

Background

Oxidation–reduction reactions are a major class of chemical reactions. An oxidation–reduction, or redox, reaction is defined as any reaction in which electrons are transferred from one substance to another. *Oxidation* occurs when a substance loses electrons. Because any electrons lost by one reactant must be transferred to another reactant, oxidation and reduction always occur together. *Reduction* occurs when a substance gains electrons.

Substances that are used to cause the oxidation or reduction of another substance are called oxidizing and reducing agents, respectively. The substance that accepts electrons in a redox reaction is called the *oxidizing agent*—by accepting electrons, it is itself reduced but it *causes* the oxidation of another substance. Similarly, the substance that loses electrons in a redox reaction is called the *reducing agent* because it causes the reduction of another substance.

The loss and gain of electrons by the reactants in a chemical reaction is not always obvious from the formulas of the reactants and products. A method based on oxidation states has been developed to identify oxidation–reduction reactions, to determine whether a substance has been oxidized or reduced, and to count the electrons that are lost or gained as a result. The *oxidation state* may be thought of as an imaginary charge on an atom in an element or compound. Oxidation states are assigned strictly for "electron bookkeeping" purposes:

1. The oxidation state of an atom in a free *element* is zero.
2. The oxidation state of an atom in a *monatomic ion* is equal to the charge on the ion.
3. The oxidation state of *fluorine* in a compound is always –1.
4. The oxidation state of *hydrogen* in a compound is +1, *except* in metal hydrides (ionic compounds with metals), where it is –1.
5. The oxidation state of *oxygen* in a compound is –2, *except* in peroxides (compounds containing O—O bonds), where it is –1.
6. The sum of the oxidation states of all the atoms in a *neutral compound* is equal to zero.
7. The sum of the oxidation states of all the atoms in a *polyatomic ion* is equal to the charge on the ion.

A reaction is classified as a redox reaction if the oxidation states of the reactants change. Oxidation is an *increase* in oxidation state (corresponding to a loss of electrons). Reduction is a *decrease* in oxidation state (corresponding to a gain of electrons). Consider the reaction of Fe^{2+} ions with chlorine (Equation 1). The reaction is identified as a redox reaction based on the change in oxidation states for iron and chlorine. Iron is oxidized—the oxidation state of iron increases from +2 to +3. Chlorine is reduced—the oxidation state of chlorine decreases from zero to –1.

$$2Fe^{2+}(aq) + Cl_2(aq) \rightarrow 2Fe^{3+}(aq) + 2Cl^{-}(aq) \qquad \textit{Equation 1}$$

Oxidation states: *(+2)* *(0)* *(+3)* *(–1)*

For every redox reaction, two separate half-reactions can be written. The *oxidation half-reaction* shows the substance that is oxidized, the product resulting from oxidation, and the number of electrons lost in the process. (The number of electrons lost is equal to the difference in oxidation states between the reactant and product.) The *reduction half-reaction* shows the substance that is reduced, the number of electrons gained in the process, and the product resulting from the reduction. The oxidation and reduction half-reactions for the redox reaction of Fe^{2+} with chlorine are shown below. The oxidation half-reaction must be multiplied by a factor of two so that the number of electrons lost by Fe^{2+} will be equal to the number of electrons gained by chlorine.

$$Fe^{2+} \rightarrow Fe^{3+} + e^{-} \qquad \textit{Oxidation half-reaction}$$

$$Cl_2 + 2e^{-} \rightarrow 2Cl^{-} \qquad \textit{Reduction half-reaction}$$

Experiment Overview

The purpose of this experiment is to investigate the reactions of Fe^{2+} and Fe^{3+} ions with oxidizing and reducing agents, respectively. The results will be analyzed to determine the change in oxidation state for each reactant, the oxidation and reduction half-reactions, and the balanced chemical equation for the overall redox reaction.

Pre-Lab Questions

1. Potassium iodate (KIO_3) is a strong oxidizing agent and will oxidize Fe^{2+} ions to Fe^{3+}. In doing so, iodate ion (IO_3^-) is reduced to elemental iodine (I_2).

   (a) Use the oxidation state rules (see the *Background* section) to assign oxidation states to the *iodine atom* in IO_3^- and I_2.

   (b) The half-reaction for the reduction of iodate is shown below. Use the *difference* in oxidation states for the iodine atom in IO_3^- and I_2 to determine the number of electrons gained in this half-reaction. *Hint:* Hydrogen ions (H^+) and water molecules (H_2O) are required to balance mass and charge.

$$IO_3^- + 6H^+ + \boxed{}\,e^- \rightarrow \tfrac{1}{2}I_2 + 3H_2O$$

2. Combine the oxidation half-reaction for Fe^{2+} (see the *Background* section) with the reduction half-reaction for iodate (Question #1b) and write the balanced equation for the overall redox reaction of Fe^{2+} with IO_3^-. *Hint:* The number of electrons on each side must cancel out.

Materials

Hydrochloric acid, HCl, 3 M, 2 mL

Iron(II) ammonium sulfate, $Fe(NH_4)_2(SO_4)_2$, 0.1 M, 5 mL

Iron(III) chloride, $FeCl_3$, 0.1 M, 7 mL

Potassium ferricyanide, $K_3Fe(CN)_6$, 0.1 M, 2 mL

Potassium thiocyanate, KSCN, 0.1 M, 1 mL

Oxidizing agents

Hydrogen peroxide, H_2O_2, 3%, 1 mL

Potassium permanganate solution, $KMnO_4$, 0.025 M, 1 mL

Sodium hypochlorite solution (household bleach), NaOCl, 5%, 1 mL

Reducing agents

Sodium bromide, NaBr, 0.2 M, 1 mL

Sodium iodide, NaI, 0.2 M, 1 mL

Sodium sulfite, Na_2SO_3, 0.2 M, 1 mL

Pineapple juice, 1 mL (optional)

Vitamin C solution, 0.2%, 1 mL

Equipment

Beral-type pipets or eyedroppers, 12

Distilled water and wash bottle

Paper towels

Reaction plate, 24-well

Toothpicks

Labels and markers

Safety Precautions

Follow all directions carefully and do not perform any unauthorized reactions. Hydrochloric acid is a corrosive liquid and toxic by ingestion or inhalation. Sodium hypochlorite solution reacts with concentrated acids to evolve poisonous chlorine gas. It is a corrosive liquid and is moderately toxic by ingestion and inhalation. Hydrogen peroxide is a strong oxidizing agent and a skin and eye irritant. Potassium ferricyanide and potassium thiocyanate solutions are toxic by ingestion and may evolve poisonous fumes upon heating or in contact with concentrated acids. 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

Part A. Reactions of Iron(II) Ions with Oxidizing Agents

1. Place a clean, 24-well reaction plate on top of a sheet of white paper, as shown in Figure 1. 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.

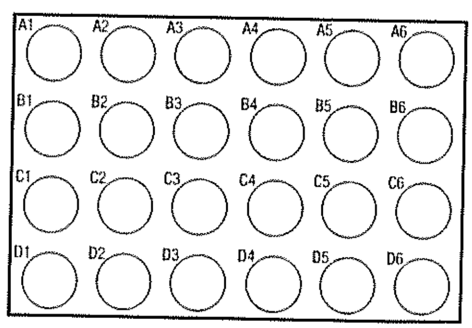

Figure 1. Layout and Numbering of a 24-Well Reaction Plate.

2. Using a clean Beral-type pipet or eyedropper for each solution, place 20 drops of iron(II) ammonium sulfate solution into well A1 and 20 drops of iron(III) chloride solution into well A2. Record the initial color of each solution in Data Table A.
3. Add 2 drops of potassium thiocyanate solution to each well A1 and A2. Record observations in Data Table A.
4. Place 20 drops of iron(II) ammonium sulfate solution into each well B1, B2, and B3.
5. Add 5 drops of 3 M hydrochloric acid solution to each well B1 and B2.
6. Using a clean pipet for each solution, add
   - 5 drops of hydrogen peroxide solution to well B1
   - 10 drops of potassium permanganate solution to well B2
   - 10 drops of sodium hypochlorite solution to well B3.
7. Use a clean toothpick to stir each solution, if needed. Record observations in Data Table A.
8. Test for the presence of iron(III) ions in wells B1, B2, and B3 by adding 5 drops of potassium thiocyanate solution to each solution. Record the final color of each test mixture in Data Table A.

Part B. Reactions of Iron(III) Ions with Reducing Agents

9. Using a clean Beral-type pipet or eyedropper for each solution, place 20 drops of iron(II) ammonium sulfate solution into well C1 and 20 drops of iron(III) chloride solution into well C2. Record the initial color of each solution in Data Table B.
10. Add 2 drops of potassium ferricyanide solution to each well C1 and C2. Record observations in Data Table B.
11. Place 20 drops of iron(III) chloride solution into each well D1–D5.
12. Add 5 drops of 3 M hydrochloric acid and 5 drops of sodium sulfite solution to well D1. Record observations in Data Table B.
13. Test for the presence of iron(II) ions in well D1 by adding 2 drops of potassium ferricyanide solution. Record the final color of the solution in Data Table B.
14. Add 5 drops of sodium bromide solution to well D2. Record observations in Data Table B, then test for the presence of iron(II) ions by adding 2 drops of potassium ferricyanide solution. Record the final color in Data Table B.
15. Add 5 drops of sodium iodide solution to well D3. Record observations in Data Table B, then test for the presence of iron(II) ions by adding 2 drops of potassium ferricyanide solution. Record the final color in Data Table B.
16. Add 10 drops of Vitamin C solution to well D4. Record observations in Data Table B, then test for the presence of iron(II) ions by adding 2 drops of potassium ferricyanide solution. Record the final color in Data Table B.

17. *(Optional)* Add 10 drops of pineapple juice to well D5. Record observations in Data Table B, then test for the presence of iron(II) ions by adding 2 drops of potassium ferricyanide solution. Record the final color in Data Table B.

18. Rinse the contents of the reaction plate down the drain with plenty of excess water. Wash the reaction plate and rinse well with distilled water.

Name: ______________________________

Class/Lab Period: ______________________

Oxidation–Reduction Survey

Data Table A. Reactions of Iron(II) Ions with Oxidizing Agents

Well	Reactants	Observations (Initial Color)	Color After Adding KSCN
A1	$Fe^{2+}(aq)$		
A2	$Fe^{3+}(aq)$		
B1	Fe^{2+} + HCl + H_2O_2		
B2	Fe^{2+} + HCl + $KMnO_4$		
B3	Fe^{2+} + NaOCl		

Data Table B. Reactions of Iron(III) Ions with Reducing Agents

Well	Reactants	Observations (Initial Color)	Color After Adding $K_3Fe(CN)_6$
C1	$Fe^{2+}(aq)$		
C2	$Fe^{3+}(aq)$		
D1	Fe^{3+} + HCl + Na_2SO_3		
D2	Fe^{3+} + NaBr		
D3	Fe^{3+} + NaI		
D4	Fe^{3+} + Vitamin C		
D5	Fe^{3+} + Pineapple Juice		

Post-Lab Questions

1. How can potassium thiocyanate be used to confirm that Fe^{2+} ions have been oxidized to Fe^{3+}?

2. Use the oxidation state rules to assign oxidation states for the indicated atoms in each oxidizing agent and its product (Part A).

Atom	Oxidizing Agent	Oxidation State	Product	Oxidation State
Mn	MnO_4^-		Mn^{2+}	
O	H_2O_2		H_2O	
Cl	OCl^-		Cl^-	

3. Fill in the blanks to show the number of electrons involved in each half-reaction.

   (a) $MnO_4^-(aq) + 8H^+(aq) + \boxed{}\ e^- \rightarrow Mn^{2+}(aq) + 4H_2O(l)$

   (b) $H_2O_2(aq) + 2H^+(aq) + \boxed{}\ e^- \rightarrow 2H_2O(l)$

   (c) $OCl^-(aq) + H_2O(l) + \boxed{}\ e^- \rightarrow Cl^-(aq) + 2OH^-(aq)$

4. Combine the oxidation half-reaction for Fe^{2+} (see the *Background* section) with the appropriate half-reaction from Question #3 and write the balanced equation for the overall redox reaction of Fe^{2+} with (a) permanganate ion, (b) hydrogen peroxide, and (c) hypochlorite ion.

5. Circle the correct choices: An *oxidizing agent* is a substance that causes the **(oxidation/reduction)** of another reactant in a redox reaction. The oxidation state of the oxidizing agent **(increases/decreases)** and the oxidizing agent itself undergoes **(oxidation/reduction)** during the reaction.

6. How can potassium ferricyanide be used to confirm that Fe^{3+} ions have been reduced to Fe^{2+}?

7. *(a)* Sulfite ion (SO_3^{2-}) is a strong reducing agent. Assign oxidation states to the sulfur atom in SO_3^{2-} and its product, sulfate ion (SO_4^{2-}).

   (b) Fill in the blank to show the number of electrons in the following half-reaction.

   $$SO_3^{2-}(aq) + H_2O(l) \rightarrow SO_4^{2-}(aq) + 2H^+(aq) + \boxed{}\, e^-$$

   (c) Write the balanced equation for the overall redox reaction of Fe^{3+} with a sulfite ion.

8. Circle the correct choices: A *reducing agent* is a substance that causes the **(oxidation/reduction)** of another substance in a redox reaction. The oxidation state of the reducing agent **(increases/decreases)** and the reducing agent itself undergoes **(oxidation/reduction)** during the reaction.

9. Based on the observations in Part B, which halide—bromide ion or iodide ion—is the stronger reducing agent? Explain.

10. Iron(II) compounds in foods are more easily absorbed by the body than iron(III) compounds. Vitamin C improves the absorption of dietary iron. Explain based on your observations in this experiment.

11. *(Optional)* Suggest a possible reason for the results obtained using pineapple juice in this experiment.