Experiment 10

Oxidation-Reduction (Redox) Reactions

Pre-Lab Assignment

Before coming to lab:

- Read the lab thoroughly.
- Answer the pre-lab questions that appear at the end of this lab exercise.

Purpose

A set of oxidation-reduction (redox) reactions will be performed to determine the relative strengths as reducing and oxidizing agents of metals and other ions. An activation series will be devised and used to predict whether a reaction will occur spontaneously or not.

Background

Oxidation-reduction (redox) reactions are a classification of chemical changes that involve the transfer of electrons. An example of a redox reaction is shown in Eqn. 1, when magnesium metal reacts with chlorine gas.

$$Mg(s) + Cl_2(g) \rightarrow MgCl_2(aq) \qquad \text{Eqn. 1}$$

At first glance, the electron exchange is not apparent as electrons are not included in the overall balanced equation. However, electrons are shown when the reaction is split into halves: an oxidized half-reaction and a reduced half-reaction, as in Eqn. 2 and Eqn. 3 respectively.

$$Mg(s) \rightarrow Mg^{2+}(aq) + 2\ e^- \qquad \text{Eqn. 2}$$

$$Cl_2(g) + 2\ e^- \rightarrow 2\ Cl^-(aq) \qquad \text{Eqn. 3}$$

Notice that in the half-reactions, the electrons are required to balance the charges. Since Mg is becoming more positive, one magnesium atom loses two electrons to become a cation and is said to be **oxidized**. Chlorine is becoming more negative, so one molecule of $Cl_2(g)$ gains two electrons to become two chloride anions and is said to be **reduced**. Both oxidation and reduction happen simultaneously in a single redox reaction. Overall, Mg is transferring two electrons to chlorine (Fig. 1).

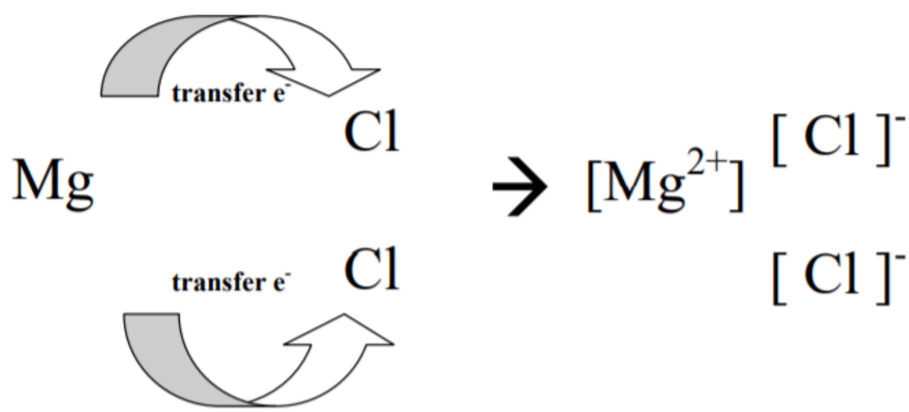


Fig. 1: Redox reaction between Mg(s) and $Cl_2(g)$ to produce $MgCl_2(aq)$

The species that loses electrons (Mg in Eqn. 2) is called the **reducing agent** and the species that gains electrons (Cl in Eqn. 3) is called the **oxidizing agent.** A reaction is said to occur spontaneously (requiring no outside intervention) if the strongest reducing agent is being oxidized and the strongest oxidizing agent is being reduced. If this is reversed, the reaction is said to be nonspontaneous and will not occur. Thus redox reactions can be considered a "competition" between reactants for electrons.

Single displacement reactions involve a pure element reacting with a compound. Successful reactions will only occur when the stronger oxidizing or reducing agent starts as the uncharged pure element. Consider Eqn. 4 and 5.

$Cu(NO_3)_2(aq) + Zn(s) \rightarrow Cu(s) + Zn(NO_3)_2(aq)$ Eqn. 4

Net Ionic Equation: $Cu^{2+}(aq) + Zn(s) \rightarrow Cu(s) + Zn^{2+}(aq)$

Oxidation Half-Reaction: $Zn(s) \rightarrow Zn^{2+}(aq) + 2\ e^-$

Reduction Half-Reaction: $Cu^{2+}(aq) + 2\ e^- \rightarrow Cu(s)$

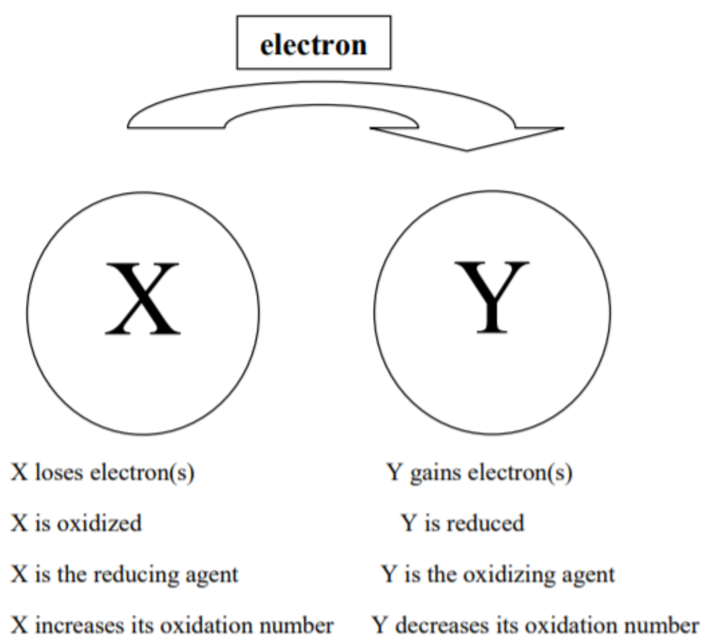
$Zn(NO_3)_2(aq) + Cu(s) \rightarrow Zn(s) + Cu(NO_3)_2(aq)$ Eqn. 5

Net Ionic Equation: $Zn^{2+}(aq) + Cu(s) \rightarrow Zn(s) + Cu^{2+}(aq)$

Oxidation Half-Reaction: $Cu(s) \rightarrow Cu^{2+}(aq) + 2\ e^-$

Reduction Half-Reaction: $Zn^{2+}(aq) + 2\ e^- \rightarrow Zn(s)$

Since Eqn. 4 and 5 are the reverse of one another, only one can occur spontaneously and the other must be nonspontaneous. In Eqn. 4, the reducing agent is Zn (was oxidized) whereas in Eqn. 5, the reducing agent is Cu. If Zn is a stronger reducing agent than Cu, then Eqn. 4 will occur spontaneously. If Cu is a stronger reducing agent than Zn, then Eqn. 5 will occur spontaneously.



Some redox reactions involving larger, more complex molecules are less obvious and oxidation numbers must be assigned to keep track of electron transfers. Oxidation numbers are not the same as ionic charges as they apply to all elements in any compound, including molecular compounds. Elements that increase (become more positive) in oxidation number are said to be oxidized and are contained in the reducing agent. Elements that decrease (become more negative) in oxidation number are said to be reduced and are contained in the oxidizing agent (Fig. 2).

Fig. 2: Summary of Redox Reactions

Rules for Assigning Oxidation Numbers

Priority starts at Rule 1 and works down. Rules higher in the list take precedence over rules appearing lower in the list.

1. The total sum of the oxidation numbers of all the atoms in a molecule must add up to the overall charge on the molecule.

Example: the oxidation numbers must sum to zero in H_2O and $CaSO_4$, and -3 in PO_4^{3-}

2. The oxidation number of a monoatomic ion is equal to the charge of the monoatomic ion.

Example: the oxidation number of S^{2-} is -2, and of Al^{3+} is +3.

3. A pure element has an oxidation number of zero.

Example: the oxidation number of Li(s) and H_2(g) is 0.

4. Metals are always positive and equal to their ionic charge.

Example: the oxidation number of Li in LiCl is +1, and of Ti in $TiCl_4$ is +4.

5. Fluorine has an oxidation number of -1.

Example: the oxidation number of F in SF_6 is -1.

6. Oxygen has an oxidation number of -2 unless present as a peroxide (-1).

Example: the oxidation number of O in MgO is -2, and in H_2O_2 is -1.

7. Hydrogen has an oxidation number of +1 when bonded to a nonmetal but -1 when bonded to a metal.

Example: the oxidation number of H in CH_4 is +1, and in LiH is -1.

8. Halogens have an oxidation number of -1.

Example: the oxidation number of Cl in NaCl is -1, and in OCl_2 is +1.

Example Problem: Assigning Oxidation Numbers

Assign oxidation numbers to all elements in MnO_4^-.

Step 1: Follow the rules down from Rule 1 until a rule applies (Rule 1).

The overall sum of oxidation numbers must equal -1.

Step 2: Follow the rules down from Rule 1 until a rule applies (Rule 8).

The oxidation number of oxygen is -2

Step 3: Solve algebraically for Mn (no rule).

$-1 = (1 \times Mn) + (4 \times -2)$

$-1 = Mn + (-8)$

Mn = +7, the oxidation number of manganese is +7.

Activity series rank elements and ions by how easily they are oxidized—in other words, how strongly they behave as reducing agents. More active elements are more easily oxidized and are stronger reducing agents, so are listed higher on the series. Less active elements are less easily oxidized and are weaker reducing agents, so are listed lower on the series.

Procedure

Part I: Reacting Metals with Acid

1. Obtain four clean small test tubes.

2. Add approximately 1 mL of 6 M HCl(aq) solution to each.

3. Obtain a small piece of magnesium, zinc, lead, and copper metal each.

4. Add one metal to each of the four test tubes in Step 2. Record your observations and determine whether or not a reaction has occurred. For any successful reaction, write the net ionic equation and determine the stronger reducing agent.

Hint: Cl^- is a spectator ion. H, when ionized, is +1 and all metals will be +2.

5. Separate the unreacted solids from their solutions and dispose of each in their appropriate waste container.

6. Create an activity series, ranking each metal as a stronger or weaker reducing agent in comparison to H.

Part II: Reacting Metals with Metal Ions

1. Obtain three clean small test tubes.

2. To each test tube, add approximately 1 mL of 0.1 M, $Pb(NO_3)_2(aq)$, $Cu(NO_3)_2(aq)$, and $Zn(NO_3)_2$ solutions. There should only be one solution in each test tube.

3. To teach test tube in Step 2, add one piece of Mg(s) metal.

4. Record your observations and determine whether or not a reaction has occurred. For any successful reaction, write the net ionic equation and determine the stronger reducing agent.

Note: some reactions are very slow. Allow at least 20 minutes before recording "no reaction".

Hint: NO_3^- is a spectator ion. All metals, when ionized, will be +2.

5. Separate the unreacted solids from their solutions and dispose of each in their appropriate waste container.

6. Wash out the test tubes with deionized water. They do not have to be completely dry.

7. Repeat Steps 1-6, following Table 1 to prepare unique pairs of metal solid and solution to react each time. There should be 12 reaction mixtures total.

8. Create an activity series, ranking Mg, Pb, Cu and Zn by increasing strength as a reducing agent.

Table 1: Reaction Mixtures for Part II

Solids	Solutions			
	$Mg(NO_3)_2(aq)$	$Pb(NO_3)_2(aq)$	$Cu(NO_3)_2(aq)$	$Zn(NO_3)_2(aq)$
Mg(s)		1	2	3
Pb(s)	4		5	6
Cu(s)	7	8		9
Zn(s)	10	11	12	

Part III: Silver Metal

1. Obtain one clean small test tube.

2. Add approximately 1 mL of 0.1 M $AgNO_3(aq)$ solution to the test tube.

3. Add a piece of Cu metal. Record your observations and determine whether or not a reaction has occurred. If so, write the net ionic equation and determine the stronger reducing agent.

Hint: NO_3^- is a spectator ion. When ionized, Cu will be +2 and Ag will be +1.

4. Based on your data from Parts I and III, rank H and Ag in your activity series from Part II.

Experiment 10—Data Sheet

Name: ____

Part I: Reacting Metals with Acid

Reaction 1: $HCl(aq) + Mg(s)$

Observations:

Did a reaction occur?

If Yes, write the net ionic equation:

Stronger Reducing Agent:

Reaction 2: $HCl(aq) + Zn(s)$

Observations:

Did a reaction occur?

If Yes, write the net ionic equation:

Stronger Reducing Agent:

Reaction 3: $HCl(aq) + Pb(s)$

Observations:

Did a reaction occur?

If Yes, write the net ionic equation:

Stronger Reducing Agent:

Reaction 4: $HCl(aq) + Cu(s)$

Observations:

Did a reaction occur?

If Yes, write the net ionic equation:

Stronger Reducing Agent:

Ranking:

Part II: Reacting Metals with Metal Ions

Reaction 1: $Mg(s) + Pb(NO_3)_2\ (aq)$

Observations:

Did a reaction occur?

If Yes, write the net ionic equation:

Stronger Reducing Agent:

Reaction 2: $Mg(s) + Cu(NO_3)_2(aq)$

Observations:

Did a reaction occur?

If Yes, write the net ionic equation:

Stronger Reducing Agent:

Reaction 3 $Mg(s) + Zn(NO_3)_2(aq)$

Observations:

Did a reaction occur?

If Yes, write the net ionic equation:

Stronger Reducing Agent:

Reaction 4: $Pb(s) + Mg(NO_3)_2(aq)$

Observations:

Did a reaction occur?

If Yes, write the net ionic equation:

Stronger Reducing Agent:

Reaction 5: $Pb(s) + Cu(NO_3)_2(aq)$

Observations:

Did a reaction occur?

If Yes, write the net ionic equation:

Stronger Reducing Agent:

Reaction 6: $Pb(s) + Zn(NO_3)_2(aq)$

Observations:

Did a reaction occur?

If Yes, write the net ionic equation:

Stronger Reducing Agent:

Reaction 7: $Cu(s) + Mg(NO_3)_2(aq)$

Observations:

Did a reaction occur?

If Yes, write the net ionic equation:

Stronger Reducing Agent:

Reaction 8: $Cu(s) + Pb(NO_3)_2(aq)$

Observations:

Did a reaction occur?

If Yes, write the net ionic equation:

Stronger Reducing Agent:

Reaction 9: $Cu(s) + Zn(NO_3)_2(aq)$

Observations:

Did a reaction occur?

If Yes, write the net ionic equation:

Stronger Reducing Agent:

Reaction 10: $Zn(s) + Mg(NO_3)_2(aq)$

Observations:

Did a reaction occur?

If Yes, write the net ionic equation:

Stronger Reducing Agent:

Reaction 11: $Zn(s) + Pb(NO_3)_2(aq)$

Observations:

Did a reaction occur?

If Yes, write the net ionic equation:

Stronger Reducing Agent:

Reaction 12: $Zn(s) + Cu(NO_3)_2(aq)$

Observations:

Did a reaction occur?

If Yes, write the net ionic equation:

Stronger Reducing Agent:

Ranking:

Part III: Silver Metal

Reaction 1: $Cu(s) + AgNO_3(aq)$

Observations:

Did a reaction occur?

If Yes, write the net ionic equation:

Stronger Reducing Agent:

Complete Activity Series (Parts I, II, and III):

Experiment 10—Post-Lab Assignment

1. Use your complete activity series to predict whether the following reactions will or will not occur. If a reaction does occur, write the net ionic equation.

a. $Zn(s) + AgNO_3(aq) \rightarrow$

b. $Ag(s) + Pb(NO_3)_2(aq) \rightarrow$

c. $Pb(s) + AgNO_3(aq) \rightarrow$

2. Imagine that the hypothetical elements A, B, C, and D form the ions A^{2+}, B^{2+}, C^{2+}, and D^{2+} respectively when they behave as reducing agents. The following reactions were observed in lab to occur or not occur. Use this information to write an activity series for A, B, C, and D, ranking them by increasing strength as reducing agents.

$$B^{2+}(aq) + A(s) \rightarrow A^{2+}(aq) + B(s)$$

$$B^{2+}(aq) + D(s) \rightarrow \text{no reaction}$$

$$A^{2+}(aq) + C(s) \rightarrow C^{2+}(aq) + A(s)$$

3. The halogens (F_2, Cl_2, Br_2, I_2) form halide ions (F^-, Cl^-, Br^-, I^-, respectively) when they behave as oxidizing agents. The following reactions were observed in lab to occur or not occur. Use this information to write an activity series for the halogens, ranking them by increasing strength as oxidizing agents.

$$F_2(g) + 2\ Cl^-(aq) \rightarrow 2\ F^-(aq) + Cl_2(l)$$

$$I_2(l) + 2\ Cl^-(aq) \rightarrow \text{no reaction}$$

$$I_2(l) + 2\ Br^-(aq) \rightarrow \text{no reaction}$$

$$Cl_2(l) + 2\ Br^-(aq) \rightarrow Br_2(l) + 2\ Cl^-(aq)$$

Experiment 10—Pre-Lab Assignment

Name: ____

1. Determine the oxidation number of each element in the substances below.

a.	MnO_2	Mn:	_________	O: _________			
b.	Fe_2O_3	Fe:	_________	O: _________			
c.	Na_2O_2	Na:	_________	O: _________			
d.	H_2	H:	_________				
e.	$K_2Cr_2O_7$	K:	_________	Cr: _________		O: _________	

2. Combustion reactions are also oxidation-reduction (redox) reactions.

a. Write the balanced chemical equation for the combustion of methane gas, CH_4.

b. Assign oxidation numbers to each element in every compound in (a).

c. Which element is being oxidized? ____

d. Which element is being reduced? ____

e. What compound is behaving as the oxidizing agent? ______________________

f. What compound is behaving as the reducing agent? ______________________

3. Solid tin will react with aqueous platinum(II) chloride to produce aqueous tin(IV) chloride and solid platinum.

a. Write the balanced molecular equation for the above reaction.

b. Write the net ionic equation for the above reaction.

c. Write the oxidation half-reaction for the above reaction.

d. Write the reduction half-reaction for the above reaction.