7.2 - Balancing Redox Reactions Using Half Reactions

pages 650-653 in Matter and Change

Reduction

Oxidant + e^- → Product

(Gain of Electrons) (Oxidation Number Decreases)

Oxidation

Reductant → Product + e^-

(Loss of Electrons) (Oxidation Number Increases)

We can use oxidation numbers to identify which reactions are redox, which element is gaining electrons, and which is losing electrons.

For example,

$2Na + Cl_2 \rightarrow 2NaCl$

It is useful to write in the oxidation of every element in every compound above the element in the equation.

Remember that the balancing coefficients in the chemical equations **do not** affect the oxidation numbers.

A chart is a useful way of organizing the changes in oxidation number for each element:

element	Initial Ox. #	Final Ox. #	Change in e^-	Oxidized or Reduced?

By looking at the table, we see:

i. oxidation numbers **did** change, so it is a redox reaction.

ii. Na increased its oxidation number from 0 to +1. Therefore, it has **lost**electrons. (**LEO**)

iii. Cl decreased its oxidation number from 0 to -1. Therefore it has **gained** electrons. (**GER**)

So, an **increase** in oxidation number indicates **oxidation**, while a **decrease** in oxidation number indicates **reduction**.

Ex 1) Consider the following reaction:

$2Mg + O_2 \rightarrow 2MgO$

Summarize the changes in oxidation number for each element, determine how many electrons has been transferred per atom, and identify what has been oxidized and what has been reduced.

element	Initial Ox. #	Final Ox. #	Change in e^-	Oxidized or Reduced?

Recall that oxidation cannot occur without reduction (and vice versa). Therefore, if one substance cannot accept the electrons a substance is giving away, they cannot be given away in the first place. That is, one **allows** for the other to occur.

- We call the substance that is oxidized the **reducing agent** because it allows another element to be reduced.
- We call the substance that is being reduced the **oxidizing agent** because it allows another element to be oxidized.

So, in our example above, since Mg was _____________ and it is the ___________________.
Likewise, O is the ___________________ because it was __________________.

Ex 2) In the chemical reaction

$N_2 + 3H_2 \rightarrow 2NH_3$

summarize the changes in oxidation numbers, determine the number of electrons transferred per atom, identify what has been oxidized or reduced, and identify the oxidizing and reducing agents.

element	Initial Ox. #	Final Ox. #	Change in e^-	Oxidized or Reduced?	Type of Agent?

**If there is no transfer of electrons, the reaction is not a REDOX reaction. Charges of atoms will remain the same in a situation like this.

7.2 - Balancing Redox Reactions with Half Reactions

Now that we can recognize redox reactions, we need to know how to balance them.

Balancing redox reactions can be tricky; therefore, there are two different methods we will look at: Half Reaction and Oxidation Number Methods

Oxidation and reduction always occur together. Each part of the reaction is called a "half-reaction" *(Table 18 and 19)*

Steps:

1. Break the reaction into **two** half reactions (an oxidation and a reduction) and remove any spectator ions from the equations (or use Table 18 and 19 to help find half reactions)

   Note - oxidation numbers may help decide what is being oxidized and what is being reduced.

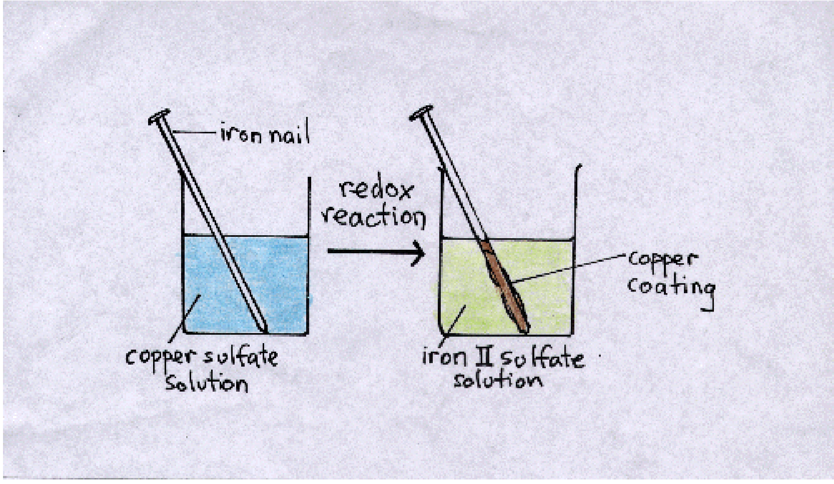
2. Balance each half-reaction separately, first by number of atoms, then by charge by adding electrons to the appropriate side of the equation.

3. Compare the number of electrons in each equation. They must be equal. If they are not equal, make them equal by multiplying everything in **one** half-reaction by a coefficient.

4. Add the two equations together and replace all spectator ions in the correct spot. Electrons do not belong here (they should cancel). Furthermore, all compounds previously broken up are written as bonded together again.

Ex) Balance the following using the half reaction method:

$$Cu_{(s)} + AgNO_{3\,(aq)} \rightarrow Cu(NO_3)_{2\,(aq)} + Ag_{(s)}$$



Ex) Balance the following using half reactions:

$Fe_{(s)} + CuSO_{4(aq)} \rightarrow Cu_{(s)} + Fe_2(SO_4)_{3\,(aq)}$

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Balancing Half Reactions in **Acidic** Solutions

When you are told that a half reaction is in an acidic solution we need to understand that the half reaction itself does not show all substances present.

Since a half reaction only explains how oxidation numbers change for a particular substance, we may need to add $\underline{H_2O}$ (since we are in a solution) and $\underline{H^+}$ (since the solution is acidic) to the equation.

Ex) Write the balanced equation for the half reaction in which nitric oxide (NO) is reduced to nitrous oxide (N_2O) in an acidic solution.

The Steps:

1. Balance all elements for mass **except** for hydrogen and oxygen.

2. Add H_2O to one side of the reaction to balance the oxygens first.

3. You will probably have a discrepancy with the number of hydrogens, so add a number of H^+ to the opposite side as H_2O to balance them out.

4. Now, add electrons to one of the sides so that each side of the reaction is electrically the same.

This is now the balanced chemical equation for this HALF-REACTION; Note, that since electrons are ______________, this process is _____________

Balancing Half Reactions in **Basic** Solutions

We will use a similar process, realizing this time that water and **OH^-** must be present.

Ex) Write the balanced equation for the oxidation half reaction in which Cl_2 is oxidized to ClO_3^- in basic solution.

The Steps:
1. Balance all elements for mass except hydrogen and oxygen.

2. Add H_2O to balance for oxygens.

3. Add H^+ to balance for hydrogens.

4. Add the same number of OH^- as H^+ ions in previous step to **both** sides of the equation.

5. Cancel off as many waters as you can (you should now have H_2O on both sides of the equation).

6. Add in electrons to make both sides equal in charge.

This is now the balanced chemical equation for this HALF-REACTION; Note, that sinc electrons are __________, _____________ has occurred in this process

7.2 Balancing Redox Reactions with Half Reactions Assignment

1. For each of these reactions, determine whether or not it is a redox reaction. If any are, identify oxidizing and reducing agents in those reactions.

a. $CaBr_2 + Pb(NO_3)_2 \rightarrow PbBr_2 + Ca(NO_3)_2$

element	Initial Ox. No		Final Ox. No.	e^- gained or lost	Oxidized or reduced	Agent
		→				
		→				

b. $P_4 + 5O_2 \rightarrow P_4O_{10}$

element	Initial Ox. No		Final Ox. No.	e^- gained or lost	Oxidized or reduced	Agent
		→				
		→				

c. $SnCl_2 + 2\ FeCl_3 \rightarrow 2\ FeCl_2 + SnCl_4$

element	Initial Ox. No		Final Ox. No.	e^- gained or lost	Oxidized or reduced	Agent
		→				
		→				

2. Break each equation into two half-reactions. Identify each half-reaction as oxidation or reduction.

a. $Cu + 2\ H^+ \rightarrow Cu^{2+} + H_2$

b. $2\ Al + 3\ S \rightarrow Al_2S_3$

3. Balance the following equations using the half-reaction method. Identify what is reduced and what is the reducing agent.

a. $Na + Br_2 \rightarrow NaBr$

b. $Zn + S \rightarrow ZnS$

c. $Au^{3+}{}_{(aq)} + Cd_{(s)} \rightarrow Au_{(s)} + Cd^{2+}{}_{(aq)}$

4. Write a balanced equation for each of the following half-reactions, and state whether it represents oxidation or reduction.

a. $HClO_2 \rightarrow Cl^-$ (acidic)

b. $Cr(OH)_3 \rightarrow CrO_4^{2-}$ (basic)

c. $H_2GeO_3 \rightarrow Ge$ (acidic)

d. $SbO_2^- \rightarrow Sb$ (basic)

5. CHALLENGE: By using **balanced** half-reaction equations from your tables work out **overall** redox equation for the reaction below in an acidic solution.

a. $Cu + NO^-_3 \rightarrow Cu^{2+} + NO$