7.3 - Balancing Redox Reactions with Oxidation Numbers

pages 644-649 in Matter and Change

I	II											III	IV	V	VI	VII
H +1																
Li +1	Be +2											B +3	C +4 +2 -4	N +5 +4 +3 +2 +1 -3	O -2	F -1
Na +1	Mg +2											Al +3	Si +4	P +5 +3 -3	S +6 +4 -2	Cl +7 +5 +3 +1 -1
K +1	Ca +2	Sc +3	Ti +4 +3	V +5 +4 +3 +2	Cr +6 +3 +2	Mn +7 +4 +3 +2	Fe +3 +2	Co +3 +2	Ni +2	Cu +2 +1	Zn +2	Ga +3 +1	Ge +4 +2	As +5 +3 -3	Se +6 +4 -2	Br +7 +5 +3 +1 -1
Rb +1	Sr +2									Ag +1	Cd +2	In +3 +1	Sn +4 +2	Sb +5 +3	Te +6 +4 -2	I +7 +5 +3 +1 -1

7.3 - Balancing Redox Reactions with Oxidation Numbers

The second balancing method for Redox reactions is the **oxidation number method.** This is very helpful for difficult equations...

For example: $Cu_{(s)} + Ag^{+}{}_{(aq)} \rightarrow Cu^{2+}{}_{(aq)} + Ag_{(s)}$

appears to be balanced, but the **charges** are not balanced.

The total electrical charge on the reactant side must equal the total electrical charge on the product side.

Therefore, like mass (Conservation of Mass), charge is conserved during a chemical reaction.

To balance the equation using oxidation numbers we will follow these steps:

1. Assign oxidation numbers to all atoms in the equation

2. Identify the atoms that are oxidized and the atoms that are reduced

3. Determine the change in the oxidation number for the atoms that are oxidized and for the atoms that are reduced

4. Make the change in oxidation numbers equal in magnitude by adjusting the coefficients of the equation

5. If necessary, use the conventional method (by mass) to balance the remainder of the equation

Ex 1) $Cu_{(s)} + HNO_{3(aq)} \text{-->} Cu(NO_3)_{2(aq)} + NO_2(g) + H_2O_{(l)}$

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Ex 2) $MnO_4^- + Fe^{2+} + H^+ \rightarrow Mn^{2+} + Fe^{3+} + H_2O$

Ex 3) $NH_3 + O_2 \rightarrow NO_2 + H_2O$

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Note - if a polyatomic ion stays intact, the oxidation numbers of its elements will not change.

Ex) $K_2Cr_2O_7$ + NaI + H_2SO_4 $\rightarrow$ $Cr_2(SO_4)_3$ + I_2 + H_2O + Na_2SO_4 + K_2SO_4

Overall Redox In Acidic and Basic Solutions

Ex) Balance the following redox reaction in an acidic solution:

$P_4 + IO_3^- \Rightarrow H_2PO_4^- + I^-$

Method 1 - Oxidation Numbers

Now we must balance for mass. Start with oxygen first by adding the H_2O molecule. Finish off by balancing the hydrogens by adding H^+.

Method 2 - Half Reactions

$P_4 + IO_3^- \Rightarrow H_2PO_4^- + I^-$

Break the reaction up into two half reactions.

1. Recall that step 1 is to balance for mass. One reaction will require you to add H_2O and H^+.

2. Balance for charge by adding the correct number of electrons in each half reaction.

3. Finally, get the number of electrons equal in both half reactions by multiplying by a coefficient.

4. Add these two half reactions up to get the overall balanced redox reaction.

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Redox reactions that are in basic solutions are similar, however we need to deal with adding OH^- as well.

Ex) Balance the equation for the reaction of the permanganate ion with the sulfite ion in a **basic** solution to give manganese dioxide (MnO_2) and the sulfate ion.

Method 1 - Oxidation Numbers

In this method, we need add H_2O on the opposite side to balance for oxygens, use H^+ to balance for charge; then add the equivalent number of OH^- ions to both sides.

Method 2 - Half Reactions

Note that each half reaction will contain oxygen, so you will need to add H_2O to both equations. (Just like in 7.2) The best way to do this is to treat this reaction as if it were in an acidic solution then neutralize the number of H^+ in the final equation with the same number of OH^- on both sides.

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7.3 – Balancing Equations with Oxidation Numbers Assignment

1. Balance the following redox reactions using the oxidation number method.

a. $SnCl_2 + HgCl_2 \text{ --> } SnCl_4 + HgCl$

b. $HNO_3 + H_2S \text{ --> } NO + S + H_2O$

c. $NaClO + H_2S \text{ --> } NaCl + H_2SO_4$

d. $MnO_4^- + H^+ + Cl^- \text{ --> } Mn^{2+} + Cl_2 + H_2O$

7.3 - Balancing Redox Reactions with Oxidation Numbers

2. Balance each of the following redox reactions in **acidic** solutions using both methods:

a) ClO_4^- + Br^- → Cl^- + BrO_3^-

i) Oxidation Numbers:

ii) Half Reactions

a) HNO_3 + Cu → NO_2 + Cu^{2+}

i) Oxidation Numbers:

ii) Half Reactions

7.3 - Balancing Redox Reactions with Oxidation Numbers

3. Balance the following redox reactions in **basic** solutions using both methods:

a) ClO_3^- + MnO_2 → Cl^- + MnO_4^-

i) Oxidation Numbers

ii) Half Reactions

b) ReO_4^- + IO^- → IO_3^- + Re

i) Oxidation Numbers

ii) Half Reactions