9.2 Balancing Redox Equations

- All redox reactions are electron transfer reactions. This means that electrons that are lost by one particle are the same electrons gained by another.
- Some redox reactions can be balanced by inspection or trial & error, but others can be very difficult to balance due to the complexity of the reaction.

Oxidation Number Method: Reactant side first then balance the product side

- The total increase in oxidation number for a particular atom/ion must equal the total decrease in oxidation number of another atom/ion.
- E.g. Hydrogen sulfide is burned and converted to sulfur dioxide.

$$H_2S_{(g)} + O_{2(g)} \rightarrow SO_{2(g)} + H_2O_{(g)}$$

First figure out oxidation numbers

Oxidation (from –2 to +4 is a change of 6 e^-)

$$\overset{+1\,-2}{H_2S_{(g)}} + \overset{0}{O_{2(g)}} \rightarrow \overset{+4\,-2}{SO_{2(g)}} + \overset{+1\,-2}{H_2O_{(g)}}$$

Reduction (from 0 to –2 is a change of 2 e^-)

+1 –2		0		+4 –2		+1 -2
$H_2S_{(g)}$	+	$O_{2(g)}$	$\rightarrow$	$SO_{2(g)}$	+	$H_2O_{(g)}$
6 e^-/S		2 e^-/O				
6 e^-/H_2S		4 e^-/O_2				

Find the simplest whole number for the coefficients for each molecule.

6 e^-/H_2S	4 e^-/O_2
×2 = 12	×3 = 12

$$2H_2S_{(g)} \quad + \quad 3O_{2(g)} \quad \rightarrow \quad SO_{2(g)} \quad + \quad H_2O_{(g)}$$

Complete by balancing the products.
Balance the atoms whose oxidation number has changed. Balance the other atoms after.

$$2H_2S_{(g)} \quad + \quad 3O_{2(g)} \quad \rightarrow \quad 2SO_{2(g)} \quad + \quad 2H_2O_{(g)}$$

- E.g. Chlorate ions and iodine react <u>in an acidic solution</u> to produce chloride ions and iodate ions.

$$ClO_3^-{}_{(aq)} + I_{2(aq)} \rightarrow Cl^-{}_{(aq)} + IO_3^-{}_{(g)}$$

First figure out oxidation numbers

+5 –2		0		–1		+5 -2
$ClO_3^-{}_{(aq)}$	+	$I_{2(aq)}$	$\rightarrow$	$Cl^-{}_{(aq)}$	+	$IO_3^-{}_{(g)}$
6 e^-/Cl		5 e^-/I				
6 e^-/ ClO_3^-		10 e^-/I_2				

Find the simplest whole number for the coefficients.

6 e^-/ ClO_3^-	10 e^-/ I_2
×5 = 30	×3 = 30

$$5ClO_3^-{}_{(aq)} \quad + \quad 3I_{2(aq)} \quad \rightarrow \quad Cl^-{}_{(aq)} \quad + \quad IO_3^-{}_{(g)}$$

Complete by balancing the products balancing the atoms whose oxidation number has changed.

$$5ClO_3^-{}_{(aq)} \quad + \quad 3I_{2(aq)} \quad \rightarrow \quad 5Cl^-{}_{(aq)} \quad + \quad 6IO_3^-{}_{(g)}$$

Problem: oxygen is not balanced 15 on left and 18 on right.

This can be fixed since it is in an acidic solution we get additional oxygen from the autoionization of water and leave the H^+ in the acidic solution. Therefore add H_2O to balance the O and H^+ to balance the H's.

$$3H_2O_{(l)} \quad + \quad 5ClO3^-{}_{(aq)} \quad + \quad 3I_{2(aq)} \quad \rightarrow \quad 5Cl^-{}_{(aq)} \quad + \quad 6IO_3^-{}_{(g)}$$

$$3H_2O_{(l)} \quad + \quad 5ClO3^-{}_{(aq)} \quad + \quad 3I_{2(aq)} \quad \rightarrow \quad 5Cl^-{}_{(aq)} \quad + \quad 6IO_3^-{}_{(g)} \quad + \quad 6H^+{}_{(aq)}$$

Summary:Copy from page 667

Step 1 Assign oxidation numbers and identify the atoms/ions whose oxidation numbers change.
Step 2 Using the change in oxidation numbers, write the number of electrons transferred per atom.
Step 3 Using the chemical formulas, determine the number of electrons transferred per reactant. (Use the formula subscripts to do this.)
Step 4 Calculate the simplest whole number coefficients for the reactants that will balance the total number of electrons transferred. Balance the reactants and products.
Step 5 Balance the O atoms using H_2O(l), and then balance the H atoms using H^+ (aq).

For basic solutions only,

Step 6 Add OH^- (aq) to both sides equal in number to the number of H^+ (aq) present.
Step 7 Combine H^+ (aq) and OH^- (aq) on the same side to form H_2O(l), and cancel the same number of H_2O(l) on both sides.

Half Reaction Method

- Balancing of redox reactions using half reactions. The reactions can occur in neutral, acidic, or basic environments. The conditions will change how you answer the question.
- E.g. Copper metal can be oxidized in an *acidic* solution to form copper (I) oxide. What is the half reaction for this process?

Step 1: write the unbalanced half reactions.
$Cu_{(s)} \rightarrow Cu_2O_{(s)}$

Step 2: Balance any atoms other than oxygen and hydrogen.
$2Cu_{(s)} \rightarrow Cu_2O_{(s)}$

Step 3: Balance oxygen by adding water.
$H_2O_{(l)} + 2Cu_{(s)} \rightarrow Cu_2O_{(s)}$

Step 4: Balance hydrogen by adding hydrogen atoms
$H_2O_{(l)} + 2Cu_{(s)} \rightarrow Cu_2O_{(s)} + 2H^+_{(aq)}$

Step 5: Balance the charges by adding electrons.
$H_2O_{(l)} + 2Cu_{(s)} \rightarrow Cu_2O_{(s)} + 2H^+_{(aq)} + 2\ e^-$

- E.g. Copper metal can be oxidized in a *basic* solution to form copper (I) oxide. What is the half reaction for this process?

$2Cu_{(s)} \rightarrow Cu_2O_{(s)}$

Complete the steps 1-4 above but add the following steps.
From Step 4:
$H_2O_{(l)} + 2Cu_{(s)} \rightarrow Cu_2O_{(s)} + 2H^+_{(aq)}$

Step 5: Adjust to basic conditions by adding the same number of hydroxide ions as hydrogen ions to both sides.
$2OH^-_{(aq)} + H_2O_{(l)} + 2Cu_{(s)} \rightarrow Cu_2O_{(s)} + 2H^+_{(aq)} + 2OH^-_{(aq)}$

Step 6: Combine hydrogen ions and hydroxide ions.
$2OH^-_{(aq)} + H_2O_{(l)} + 2Cu_{(s)} \rightarrow Cu_2O_{(s)} + 2H_2O_{(l)}$

Step7: Cancel out things on both sides.(specifically water)
$2OH^-_{(aq)} + 2Cu_{(s)} \rightarrow Cu_2O_{(s)} + H_2O_{(l)}$

Step 8: Balance the charges by adding electrons.
$2OH^-_{(aq)} + 2Cu_{(s)} \rightarrow Cu_2O_{(s)} + H_2O_{(l)} + 2\ e^-$

Balancing Redox Equations Using Half Reaction Equations

- E.g. In a chemical analysis, a solution of dichromate ions is reacted with an acidic solution of iron (II) ions. The products formed are iron (III) and chromium (III) ions as shown in the equation. Balance the redox equation.
$Fe^{2+}_{(aq)} + Cr_2O_7^{2-}_{(aq)} \rightarrow Fe^{3+}_{(aq)} + Cr^{3-}_{(aq)}$

Split the equation into half reactions Treat each half separately and balance.

$Fe^{2+}_{(aq)} \rightarrow Fe^{3+}_{(aq)}$ $Cr_2O_7^{2-}_{(aq)} \rightarrow Cr^{3-}_{(aq)}$

<table>
<tr><td>$Fe^{2+}_{(aq)} \rightarrow Fe^{3+}_{(aq)} + e^-$</td><td>$Cr_2O_7^{2-}_{(aq)} \rightarrow Cr^{3-}_{(aq)}$</td></tr>
<tr><td></td><td>$Cr_2O_7^{2-}_{(aq)} \rightarrow 2Cr^{3-}_{(aq)}$</td></tr>
<tr><td></td><td>$Cr_2O_7^{2-}_{(aq)} \rightarrow 2Cr^{3-}_{(aq)} + 7H_2O_{(l)}$</td></tr>
<tr><td></td><td>$14H^+_{(aq)} + Cr_2O_7^{2-}_{(aq)} \rightarrow 2Cr^{3-}_{(aq)} + 7H_2O_{(l)}$</td></tr>
<tr><td></td><td>$18\ e^- + 14H^+_{(aq)} + Cr_2O_7^{2-}_{(aq)} \rightarrow 2Cr^{3-}_{(aq)} + 7H_2O_{(l)}$</td></tr>
<tr><td colspan="2">number of electrons must be equal</td></tr>
<tr><td>$18[Fe^{2+}_{(aq)} \rightarrow Fe^{3+}_{(aq)} + e^-]$</td><td></td></tr>
<tr><td>$18Fe^{2+}_{(aq)} \rightarrow 18Fe^{3+}_{(aq)} + 18e^-$</td><td>$18\ e^- + 14H^+_{(aq)} + Cr_2O_7^{2-}_{(aq)} \rightarrow 2Cr^{3-}_{(aq)} + 7H_2O_{(l)}$</td></tr>
<tr><td colspan="2">Put the two reactions together and cancel out thing on both sides of the equation.</td></tr>
<tr><td colspan="2">$18Fe^{2+}_{(aq)} \rightarrow 18Fe^{3+}_{(aq)} + 18e^-$</td></tr>
<tr><td colspan="2">$18\ e^- + 14H^+_{(aq)} + Cr_2O_7^{2-}_{(aq)} \rightarrow 2Cr^{3-}_{(aq)} + 7H_2O_{(l)}$</td></tr>
<tr><td colspan="2">$18Fe^{2+}_{(aq)} + 14H^+_{(aq)} + Cr_2O_7^{2-}_{(aq} \rightarrow 2Cr^{3-}_{(aq)} + 7H_2O_{(l)} + 18Fe^{3+}_{(aq)}$</td></tr>
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