Chapter 9: Introduction of electrochemistry

You will be able to:

1 - Determine oxidation number
2 - Identify reduction and oxidation reaction
3 - Predict spontaneous reaction
4 - Balance reactions in acidic or basic solution by:
    Half– cell reaction
    Oxidation number method

Balancing Half-Reaction

- half-reactions must be balanced for mass & charge
- steps to balancing a half-reaction:
  1. balance all major **atoms other than O & H**
  2. balance **oxygens** by adding water (H_2O) molecules
  3. assume solutions are acidic and balance **hydrogens** by adding H^+
  4. balance the **charge** by adding electrons (e-)

ex. Balance the following half-reaction:

$$Cr_2O_7^{2-} \rightarrow Cr^{3+}$$

- in basic solutions . . . balance the equation as if it were acidic, then convert to basic by adding **equal numbers of hydroxide ions (OH^-) to both sides** of the equation and **cancelling out the H^+ as water**

ex. Balance the following half-reaction:

$$MnO_4^- \rightarrow MnO_2$$

Practice: Balance the following half-reactions:

1. $S_2O_8^{2-} \rightarrow HSO_4^-$ (acidic solution)

2. $N_2H_4 \rightarrow N_2$ (basic solution)

Balancing Redox reaction Using Half Reaction

An **overall redox equation** can be obtained by:

1. breaking into separate reduction and oxidation half-reactions
2. balancing each half
3. adding the two half reactions together once the number of e- lost in oxidation are balanced by the e- gained by reduction

ex. Balance the following redox reaction:

$$ClO_4^- + I_2 \rightarrow Cl^- + IO_3^-$$

Balanced redox equations **do not show e-** and the **number of atoms and the total charge** are balanced on both sides of the equation.

In **basic solutions**, the final equation can be converted by adding equal numbers of hydroxide molecules to both sides of the the equation and cancelling out the water molecules

$$5ClO_4^- + 4I_2 + 4H_2O \rightarrow 5Cl^- + 8IO_3^- + 8H^+$$

- for basic solution, add $8OH^-$ to both sides

$$5ClO_4^- + 4I_2 + 4H_2O \rightarrow 5Cl^- + 8IO_3^- + 8H^+$$

- cancel out water

Balancing Redox reaction Using Oxidation Number Method

The following steps are an effective problem-solving approach:

1. Write the unbalanced chemical equation from the given information. Determine the oxidation numbers for each element in the equation and identify the elements for which the oxidation numbers change.
2. Adjust the values of the coefficients to balance the electrons transferred.
3. Balance the rest of the equation by inspection. If necessary, balance oxygen by adding water.
4. If necessary, balance hydrogen by adding $H^+(aq)$ and/or $OH^-(aq)$.
5. Check your answer.
6. Write the balanced equation.

Examples: Balance using oxidation number method

Concentrated nitric acid, $HNO_3(aq)$, is very reactive: it oxidizes copper metal to produce toxic nitrogen dioxide gas, dissolved copper(II) nitrate, and water (**Figure 2**). Use the oxidation numbers method to write a balanced chemical equation for this reaction.

Figure 2 Nitrogen dioxide is a toxic reddish-brown gas produced when concentrated nitric acid oxidizes copper metal.

$$4\,HNO_3(aq) + Cu(s) \rightarrow 2\,NO_2(g) + Cu(NO_3)_2(aq) + 2\,H_2O(l)$$

Examples: Balance in acidic solution

To determine the iron content of an ore sample, you can perform a titration using an acidified solution of potassium permanganate, $KMnO_4(aq)$. Before the titration begins, all of the iron in the sample is converted to $Fe^{2+}(aq)$ ions. The net ionic equation that occurs during the titration is

$$MnO_4^-(aq) + Fe^{2+}(aq) \rightarrow Mn^{2+}(aq) + Fe^{3+}(aq)$$ WEB LINK

Balance the equation for this reaction.

$$MnO_4^-(aq) + 5\,Fe^{2+}(aq) + 8\,H^+(aq) \rightarrow Mn^{2+}(aq) + 5\,Fe^{3+}(aq) + 4\,H_2O(l)$$

Examples: Balance in acidic solution

The process for balancing equations in basic solutions is similar to balancing equations in acidic solutions. The only difference is that you add hydroxide ions at the end to account for the reaction taking place in a basic solution. For example, iodate ions, $IO_3^-(aq)$, react with oxalate ions, $C_2O_4^{2-}(aq)$, in a basic solution to produce carbon dioxide gas and aqueous iodide ions.

$$IO_3^-(aq) + 3\ C_2O_4^{2-}(aq) + 3\ H_2O(l) \rightarrow I^-(aq) + 6\ CO_2(g) + 6\ OH^-(aq)$$

Practice: Use both Half-Reaction & Oxidation number Method

a) $MnO_4^- + Fe^{2+} \rightarrow Mn^{2+} + Fe^{3+}$ (acidic)

b) $S^{2-} + ClO_3^- \rightarrow Cl^- + S$ (basic)

c) $Zn + As_2O_3 \rightarrow AsH_3 + Zn^{2+}$ (basic)

d) $P_4 \rightarrow H_2PO_2^- + PH_3$ (acidic)

Disproportionation: Same particle undergoes both oxidation and reduction

Balance the following redox reaction (acidic)

$Mn^{2+} + HBiO_3 \rightarrow Bi^{3+} + MnO_4^-$

Homework

- Complete Career Planning Final Assignment
- **Read Read Read** Chapter 9: pg608 – 616
- Do page 617. #1 – 9

Final is coming!

You guys can do it!