Chapter 21 – Principles of Reactivity: Electron Transfer Reactions

21.1 Oxidation/Reduction Reactions

- Terms
  - Oxidation – LEO THE LION GOES GER; Loses electrons. More positive
  - Reduction – Gain electrons. More negative
  - Reducing Agent – Oxidized – Electron donor
  - Oxidizing Agent – Reduced – Electron acceptor
  - Oxidation Number – know how to assign oxidation numbers
  - Battery
  - Electrochemical cell
  - Electrolysis
  - Galvanic cell - SPONTANEOUS
  - Voltaic cell – SPONTANEOUS
  - Electrolytic cell - NONSPONTANEOUS
- Balancing Equations for Oxidation –Reduction Reactions – Example 21.1 and Exercise 21.1
  - Half-reactions
  - Steps to balance an equation for oxidation-reduction reaction
    1) Is the reaction REDOX?
    2) Split into half reactions.
       a) Oxidation - electrons as a product (RIGHT SIDE)
       b) Reduction – electrons as a reactant (LEFT SIDE)
    3) Balance half-reaction for mass (number of atoms)
    4) Balance half-reactions for charge (number of electrons)
    5) Multiply each half-reaction by a factor to balance electrons.
    6) Add half reactions to produce overall balanced equation
    7) Check for mass (atoms) and charge (electrons) balance
- Balancing Equations for Oxidation and Reduction Reactions in Acid Solution - Examples 21.2 and 21.3, Exercise 21.2
  1) Is the reaction REDOX?
  2) Split into half reactions.
     a) Oxidation - electrons as a product (RIGHT SIDE)
     b) Reduction – electrons as a reactant (LEFT SIDE)
  3) Balance half-reaction for mass (number of atoms)
     a) Use H_2O to balance O
     b) Use H^1 to balance H.
  4) Balance half-reactions for charge (number of electrons)
  5) Multiply each half-reaction by a factor to balance electrons.
  6) Add half reactions to produce overall balanced equation
  7) Check for mass (atoms) and charge (electrons) balance
- Balancing Equations for Oxidation and Reduction Reactions in Basic Solution - Examples 21.4 and 21.5, Exercise 21.3
  1) Is the reaction REDOX?
  2) Split into half reactions.
     a) Oxidation - electrons as a product (RIGHT SIDE)
     b) Reduction – electrons as a reactant (LEFT SIDE)
  3) Balance half-reaction for mass (number of atoms)
     a) Use 2 OH^- to balance O
     b) Use H_2O to balance O *ON OPPOSITE SIDE*.
  4) Balance half-reactions for charge (number of electrons)
  5) Multiply each half-reaction by a factor to balance electrons.
  6) Add half reactions to produce overall balanced equation
  7) Check for mass (atoms) and charge (electrons) balance

- Balancing Equations for Oxidation and Reduction Reactions in Basic Solution – Dr. Broderick's Method
    1) Is the reaction REDOX?
    2) Split into half reactions.
        a) Oxidation - electrons as a product (RIGHT SIDE)
        b) Reduction – electrons as a reactant (LEFT SIDE)
    3) Balance half-reaction for mass (number of atoms)
        a) Use H_2O to balance O
        b) Use H^1 to balance H.
    4) Balance half-reactions for charge (number of electrons)
    5) Multiply each half-reaction by a factor to balance electrons.
    6) Add half reactions to produce overall balanced equation
    7) Check for mass (atoms) and charge (electrons) balance
    8) Add OH^- to *BOTH SIDES* to cancel all H^+.

- Problem Solving Tips and Ideas – 21.1
    - Never add O^{2-}, O, or O_2 to balance oxygen
    - Never add H or H_2 to balance hydrogen
    - To balance Oxygen
        - Add H_2O in acid solution
        - Add OH^- in base solution
    - Acid Solution - H^+/H_2O
    - Base Solution – OH^-/H_2O
    - Write all charges!

21.2 Chemical Change Leading to an Electric Current, Example 21.6 and Exercise 21.4

- Terms
    - Electrode – Electrons flow from Anode to Cathode
    - Salt Bridge
    - Anode – Oxidation
    - Cathode – Reduction
    - *Alphabetical Order*

21.3 Electrochemical Cells and Potentials, Example 21.6 and Exercise 21.4

- Terms
    - Electromotive force, emf
    - $\text{Electric Work} = \text{Charge} \times \text{Potential Energy Difference}$
    - Coulomb, (C)
    - $\text{Volt} = \dfrac{1\ \text{joule}}{1\ \text{coulomb}}$
    - $\text{Work (joule)} = 1\ \text{volt} \times 1\ \text{coulomb}$
    - Standard conditions
    - Standard electrode potential, E°

- E° and G° Example 21.7 and Exercise 21.5
    - $\Delta G^\circ_{rxn} = -nFE^\circ$ *All product favored electron transfer reactions have a positive* E°. *Reactant favored reactions have a negative* E°
    - F - Faraday constant 9.6485309×10^4 J/V mol

- Calculating the Potential E° of an Electrochemical Cell Example 21.8 and Exercise 21.6
    - Standard hydrogen electrode

21.4 Using Standard Potentials Examples 21.9 and 21.10 and Exercises 21.7 and 21.8

- Terms and Concepts

- E°_{net} is positive (+) for a spontaneous (product favored) reaction. ΔG°_{rxn} is negative (look at equation)
- E°_{net} is negative (-) for a spontaneous (product favored) reaction. ΔG°_{rxn} is positive (look at equation)

- Table of Standard Reduction Potentials – STANDARD CONDITIONS
  - Reactions are written as reduction (oxidized form + electrons --> reduced form)
  - Reactions written as oxidation (reduced form --> oxidized form + electrons) the sign of E° is reversed.
  - The more positive the E°, the stronger the oxidizing agent.
  - The more positive the E°, the more the species wants to be reduced.
  - The more positive the E°, the weaker the reducing agent.
  - The more negative the E°, the stronger the reducing agent.
  - The more negative the E°, the more the species wants to be oxidized.
  - The more negative the E°, the weaker the oxidizing agent.

21.5 Electrochemical Cells at Nonstandard Conditions Examples 21.9 and 21.10 and Exercises 21.7 and 21.8

- Terms and Concepts
  - The Nernst Equation allows for determination of the potential produced by a cell under nonstandard conditions.
  - The Nernst Equation at E=0 (equilibrium) allows for determination of an equilibrium constant for a reaction from a calculation or measurements of E.

- The Nernst Equation
  - $E = E^{\circ} - \frac{0.0257}{n} \ln Q \text{ at } 25°\text{C}$
  - $E = E^{\circ} - \frac{0.0592}{n} \log Q \text{ at } 25°\text{C}$

- E° and the Equilibrium Constant
  - $\ln K = \frac{nE^{\circ}}{0.0257}$ at 25°C

21.6 Batteries and Fuel Cells

- Terms and Concepts
  - Primary Batteries
    - Dry Cell Battery
    - Alkaline Battery
    - Mercury Battery
    - Lithium Battery
  - Secondary Batteries
    - Storage Batteries
    - Rechargeable Batteries
    - Lead Storage
    - Nickel-Cadmium Batteries
  - Fuel Cells

21.7 Corrosion: Redox Reactions in the Environment

- Terms and Concepts
  - Corrosion
  - Anodic inhibition
  - Cathodic protection
  - Galvanized iron
  - Sacrificial anode

21.8 Electrolysis: Chemical Change from Electric Energy Example 21.13 and Exercises 21.11

- Terms and Concepts
  - Electrolysis

21.9 Counting Electrons Examples 21.14 and 21.15 and Exercise 21.12

- Terms and Concepts
  - Current
    - $Current, I(amperes, A) = \frac{charge(coulombs, C)}{time(seconds, s)}$
  - Ampere

21.10 The Commercial Production of Chemicals by Electrochemical Methods

- Terms and Concepts
  - Aluminum
  - Chlorine and Sodium Hydroxide