Chapter 20 Notes
Oxidation-Reduction Reactions

20.1 The Meaning of Oxidation and Reduction
- What are Oxidation and Reduction?
  - Oxygen and Redox
    - **KEY**: The substance gaining O is oxidized, while the substance losing O is reduced
    - **Oxidation-Reduction Reactions**: Reactions with a substance being oxidized and another substance being reduced
    - **Redox Reactions**: Another name for oxidation-reduction reactions
  - Electron Shift in Redox Reactions
    - **Oxidation**: The complete or partial loss of electrons or gain of oxygen
    - **Reduction**: The complete or partial gain of electrons or loss of oxygen
  - Redox Reactions That Form Ions
    - **KEY**: Losing electrons is oxidation, Gaining electrons is reduction
    - **Reducing Agent**: The substance that loses electrons
    - **Oxidizing Agent**: The substance that gains electrons

Conceptual Problem 20.1 – Identifying Oxidized and Reduced Reactants
Determine what is oxidized and what is reduced, as well as identifying the oxidizing agent and reducing agent for:

$$2\text{AgNO}_3(\text{aq}) + \text{Cu(s)} \rightarrow \text{Cu(NO}_3)^2(\text{aq}) + 2\text{Ag(s)}$$

Writing the equation as an ionic equation, you can determine which compounds are gaining/losing charge:

$$2\text{Ag}^+(\text{aq}) + 2\text{NO}_3^-(\text{aq}) + \text{Cu(s)} \rightarrow \text{Cu}^{2+} + 2\text{NO}_3^-(\text{aq}) + 2\text{Ag(s)}$$

Copper is losing electrons = It is being oxidized, which makes it the reducing agent
Silver is gaining electrons = It is being reduced, which makes it the oxidizing agent

- Redox With Covalent Compounds
  - Reactions between metals and nonmetals are the easiest redox reactions because the metal loses electrons and the nonmetal gains electrons
  - Reactions between nonmetals that share electrons can result in a redistribution of electrons which makes them redox reactions as well
    - Hydrogen and Oxygen reacting to form Water is also a redox reaction
      - The electrons shift away from the hydrogen (oxidized)
      - This makes hydrogen the reducing agent
      - The electrons shift toward the oxygen (reduced).
      - This makes the oxygen the oxidizing agent

- Corrosion
  - **KEY**: Iron, a common construction metal often used in the form of the alloy steel, corrodes by being oxidized to ions of iron by oxygen
  - Resistance to Corrosion
    - Some metals are called “noble metals” because they are very resistant to losing their electrons by corrosion. Gold and Platinum are examples of these metals
    - Other metals are easily corroded by losing their electrons easily.
      - Aluminum forms tightly packed oxide coating that protects the rest of the aluminum atoms from corrosion
      - Iron forms a loosely packed oxide coating that does not protect the rest of the iron atoms from corrosion
  - Controlling Corrosion
    - To prevent corrosion, a metal surface may be coated with oil, paint, plastic, or another metal. These things protect the surface from air and water which prevents corrosion, but if that coating is scratched or worn away, the exposed metal will begin to corrode
    - Another method is when a metal is “sacrificed” or allowed to corrode, in order to “save” the second metal. To protect an iron object, a piece of magnesium (or another active metal) may be placed in electrical contact with the iron. If the iron begins to corrode, the magnesium will transfer its electrons to prevent the iron from oxidizing.
20.2 Oxidation Numbers

- Assigning Oxidation Numbers
  - **Oxidation Number** = A positive or negative number assigned to an atom to indicate its degree of oxidation or reduction
  - **KEY** = As a general rule, a bonded atom’s oxidation number is the charge that it would have if the electrons in the bond were assigned to the atom of the more electronegative element
  - **Rules for Assigning Oxidation Numbers**
    1. Monatomic Ions are equal to the magnitude and sign of its ionic charge
    2. Hydrogen in a compound is +1, unless in a metal hydride (NaH), which makes it –1
    3. Oxygen is –2, unless it is a peroxide, which makes it –1
    4. Oxidation number of any element in its natural state is 0
    5. In a neutral compound, the sum of the oxidation numbers must be equal to 0
    6. For a polyatomic ion, the sum of the oxidation numbers must equal the charge of the ion

**Conceptual Problem 20.2 – Assigning Oxidation Numbers to Atoms**

What is the oxidation number of each kind of atom in the following ions and compounds?

- SO₂
  - Oxygens are -2, so the sulfur must be +4 to maintain neutral charge
- CO₃²⁻
  - Oxygens are -2, so the carbon must be +4 to get to the -2 charge
- Na₂SO₄
  - Sodium has an ionic charge of +1, oxygens are -2, and sulfur must be +6
- (NH₄)₂S
  - Hydrogen must be +1, sulfur must be -2, which makes each nitrogen -3

- Oxidation-Number Changes in Chemical Reactions
  - **KEY** = An increase in the oxidation number of an atom or ion indicates oxidation
  - **KEY** = A decrease in the oxidation number of an atom or ion indicates reduction

**Conceptual Problem 20.3 – Identifying Oxidized and Reduced Atoms**

Use the changes in oxidation number to identify which atoms are oxidized and reduced in the following reaction:

\[ \text{Cl}_2(g) + 2\text{HBr}(aq) \rightarrow \text{Br}_2(g) + 2\text{HCl}(aq) \]

Chlorine gas and bromine gas are in their natural states, making their O.N. = 0

Hydrogen is +1 on both sides of the reaction, with Br being –1 on the left and Cl being –1 on the right

Cl is changing from 0 to –1 making it reduced

Br is changing from –1 to 0 making it oxidized

20.3 Balancing Redox Equations

- Identifying Redox Reactions
  - **KEY** = If the oxidation number of an element in a reacting species changes, that element has undergone oxidation or reduction, therefore, the reaction as a whole must be a redox reaction

**Conceptual Problem 20.4 – Identifying Redox Reactions**

Use oxidation numbers to identify whether each reaction is a redox reaction or another type:

\[ \text{Cl}_2(g) + 2\text{NaBr}(aq) \rightarrow \text{Br}_2(g) + 2\text{NaCl}(aq) \]

Reactants  Cl = 0, Na = +1, Br = –1,  Products  Br = 0, Na = +1, Cl = –1

Because Cl is being reduced, and Br is being oxidized, this reaction is a Redox Reaction

\[ 2\text{NaOH}(aq) + \text{H}_2\text{SO}_4(aq) \rightarrow \text{Na}_2\text{SO}_4(aq) + 2\text{H}_2\text{O}(l) \]

Reactants  Na = +1, O = –2, H = +1, S = +6  Products  Na = +1, S = +6, O = –2, H = +1

Since all atoms have the same charge from beginning to end, this is NOT a Redox Reaction
Two Ways to Balance Redox Equations
  o Using Oxidation-Number Changes
    • **Oxidation-Number-Change-Method** = Balancing a redox equation by comparing the increases and decreases in oxidation numbers during a reaction
      1) Assign O.N. to all the atoms in the equation and write them above the atoms
      2) Identify which atoms are oxidized and which are reduced
      3) Use one bracketing line to connect the atoms that undergo oxidation and another such line to connect those that undergo reduction
        o KEY = In a balanced redox equation, the total increase in the oxidation number of the species oxidized must be balanced by the total decrease in the oxidation number of the species being reduced
      4) Make the total increase in oxidation number equal to the total decrease in oxidation number by using appropriate coefficients
      5) Finally, make sure that the equation is balanced for both atoms and charge

Conceptual Problem 20.5 – Balancing Redox Equations by Oxidation-Number Change
Balance the following redox equation by using the oxidation-number-change method:

\[ K_2Cr_2O_7(aq) + H_2O(l) + S(s) \rightarrow KOH(aq) + Cr_2O_3(s) + SO_2(g) \]

**Step 1 = Assign Oxidation Numbers**

Reactants K = +1, O = −2, Cr = +6, H = +1, S = 0

Products K = +1, O = −2, Cr = +3, S = +4

Cr is being reduced because it is gaining electrons, S is being oxidized because it is losing electrons

**Step 2 = Identify which atoms are oxidized and which are reduced**
Cr \textsuperscript{6+} with Cr \textsuperscript{3+} = 3 electrons gained, S \textsuperscript{0} with S \textsuperscript{4+} = 4 electrons lost

**Step 3 = Connect atoms that change oxidation numbers**

**Step 4 = Make oxidation and reduction numbers equal by using coefficients**

12 is the first common multiple of 3 and 4, so there must be 4 Cr and 3 S to transfer electrons

\[ 2K_2Cr_2O_7(aq) + 3H_2O(l) + 3S(s) \rightarrow 4KOH(aq) + 2Cr_2O_3(s) + 3SO_2(g) \]

**Step 5 = Make sure the rest of the equation is balanced**

Need 4 K which requires 4 H. Oxygens balance with a 4 on KOH and 2 on H\textsubscript{2}O

\[ 2K_2Cr_2O_7(aq) + 2H_2O(l) + 3S(s) \rightarrow 4KOH(aq) + 2Cr_2O_3(s) + 3SO_2(g) \]

Using Half-Reactions
  • **Half-Reaction** = An equation showing just the oxidation or just the reduction that takes place in a redox reaction
  • **Half-Reaction Method** = You write and balance the oxidation and reduction half-reactions separately before combining them into a balanced redox equation

1) Write the unbalanced equation in ionic form
2) Write separate half-reactions for the oxidation and reduction processes
3) Balance the atoms in the half-reactions
4) Add enough electrons to one side of each half-reaction to balance the charges
5) Multiply each half-reaction by an appropriate number to make the numbers of electrons equal in both
6) Add the balanced half-reactions to show an overall equation
7) Add the spectator ions and balance the equation
Choosing a Balancing Method

- Some can be balanced the same way as all other reactions
  - This method usually works well if the oxidized and reduced agent only appears once on each side of the chemical equation and no acids or bases are involved
- Some can be balanced by the half-reaction method
  - This method usually works well if the same element being both oxidized and reduced or are being done in acidic or basic solutions