Chemical Reactions

Chapter 9: Types of Reactions, Oxidation and Reduction

Chemical reactions involve the breaking and forming of chemical bonds, leading to new substances. They can be classified in different ways.

9.1 Types of Reactions

1. **Combination (Synthesis) Reaction:** Two or more reactants combine to form a single, more complex product.
    - A + B ➡️ AB
    - Example: 2Mg(s) + O2(g) ➡️ 2MgO(s)
2. **Decomposition Reaction:** A single compound breaks down into two or more simpler substances. Requires energy input (heat, light, electricity).
    - AB ➡️ A + B
    - Example: CaCO3(s) ↔ {{heat}} CaO(s) + CO2(g)
3. **Displacement (Single Displacement) Reaction:** A more reactive element displaces a less reactive element from its compound.
    - A + BC ➡️ AC + B
    - Example: Zn(s) + CuSO4(aq) ➡️ ZnSO4(aq) + Cu(s) (Zinc is more reactive than copper)
4. **Double Decomposition (Double Displacement/Metathesis) Reaction:** The positive ions (cations) and negative ions (anions) of two different ionic compounds swap places, forming two new compounds. Often results in the formation of a precipitate, gas, or water.
    - AB + CD ➡️ AD + CB
    - Example (Precipitation): AgNO3(aq) + NaCl(aq) ➡️ AgCl(s) + NaNO3(aq)
    - Example (Neutralization): HCl(aq) + NaOH(aq) ➡️ NaCl(aq) + H2O(l)
    - Example (Gas formation): Na2CO3(s) + 2HCl(aq) ➡️ 2NaCl(aq) + H2O(l) + CO2(g)
5. **Catalytic Reactions:** Reactions that are sped up (or sometimes slowed down) by the presence of a **catalyst**.
    - **Characteristics of a Catalyst:**
        - Increases the rate of a chemical reaction without being consumed in the reaction.
        - Can be recovered chemically unchanged at the end of the reaction.
        - A small amount of catalyst can affect a large amount of reactants.
        - Does not change the position of equilibrium (for reversible reactions); it only helps equilibrium be reached faster.
        - Provides an alternative reaction pathway with lower activation energy.

    - **Types of Catalysts:**
        - **Homogeneous Catalyst:** Catalyst is in the same phase as the reactants (e.g., Fe^{2+} ions in the reaction between peroxodisulphate and iodide ions).
        - **Heterogeneous Catalyst:** Catalyst is in a different phase from the reactants (e.g., solid iron in the Haber process for ammonia synthesis, solid nickel in hydrogenation of oils).
6. **Reversible Reactions:** Reactions that can proceed in both the forward and reverse directions simultaneously. Indicated by a double arrow (↔).
    - Example: N2(g) + 3H2(g) ↔ 2NH3(g) (Haber Process)
    - Eventually, a reversible reaction reaches a state of **chemical equilibrium** where the rates of the forward and reverse reactions are equal.
7. **Thermal Dissociation:** A type of decomposition reaction where a compound breaks down reversibly into simpler substances when heated.
    - Example: NH4Cl(s) ↔ NH3(g) + HCl(g)

9.2 Oxidation and Reduction Reactions (Redox Reactions)

Redox reactions are reactions that involve a change in the oxidation states (or transfer of electrons) of atoms.

- **Definitions of Oxidation and Reduction:**
    - **Oxidation (OIL RIG - Oxidation Is Loss, Reduction Is Gain):**
        - Gain of oxygen.
        - Loss of hydrogen.
        - **Loss of electrons.**
        - Increase in oxidation number.
    - **Reduction:**
        - Loss of oxygen.
        - Gain of hydrogen.
        - **Gain of electrons.**
        - Decrease in oxidation number.
- **Electron Transfer in Redox Reactions:**
    - In an actual redox reaction, electrons are transferred from one species to another.
    - The species that loses electrons is **oxidized**.
    - The species that gains electrons is **reduced**.
    - These two processes always occur simultaneously; you cannot have one without the other.
- **Oxidation Number (or Oxidation State):** A number assigned to an element in a compound or ion that represents the charge it would have if electrons were transferred completely.
    - **Rules for Assigning Oxidation Numbers:**
        1. Elements in their uncombined state have an oxidation number of 0 (e.g., O2, Na, Cl2).
        2. For a monatomic ion, the oxidation number is equal to its charge (e.g., Na^+ is +1, Cl^- is -1).
        3. Oxygen usually has an oxidation number of -2 (except in peroxides where it's -1, or with fluorine where it's +2).
        4. Hydrogen usually has an oxidation number of +1 (except in metal hydrides where it's -1, e.g., NaH).

5. The sum of oxidation numbers in a neutral compound is 0.
6. The sum of oxidation numbers in a polyatomic ion equals the charge of the ion.

- **Oxidizing and Reducing Agents:**
  - **Oxidizing Agent (Oxidant):** The substance that **causes oxidation** (by gaining electrons itself). The oxidizing agent is **reduced**.
    - Example: O_2, $KMnO_4$, $K_2Cr_2O_7$, Cl_2, HNO_3, concentrated H_2SO_4.
  - **Reducing Agent (Reductant):** The substance that **causes reduction** (by losing electrons itself). The reducing agent is **oxidized**.
    - Example: H_2, C, CO, active metals (Na, Mg, Zn), H_2S, SO_2.
- **Electrolytic Oxidation and Reduction:** Redox reactions can be driven by electric current in a process called **electrolysis**.
  - **Oxidation occurs at the anode (positive electrode).**
  - **Reduction occurs at the cathode (negative electrode).**
- **Balancing Redox Equations:** Can be balanced using the **oxidation number method** or the **half-reaction (ion-electron) method**. For secondary school, often focus on identifying changes in oxidation numbers or simple electron transfers.