

8.9 Oxidation–Reduction Reactions

- The reactions in which electrons are transferred from one reactant to the other are called **oxidation–reduction** reactions.

Atoms that **lose electrons** are being **oxidized**.

- The reaction is referred to as **oxidation**.

Atoms that **gain electrons** are being **reduced**.

- The reaction is referred to as **reduction**.

The following rules are in order of priority:

- Free elements have an oxidation state = 0.
 - Na = 0 and Cl₂ = 0 in 2 Na(s) + Cl₂(g)
- Monatomic ions have an oxidation state equal to their charge.
 - Na = +1 and Cl = -1 in NaCl
- (a) The sum of the oxidation states of all the atoms in a compound is 0.
 - Na = +1 and Cl = -1 in NaCl, (+1) + (-1) = 0
 (b) The sum of the oxidation states of all the atoms in a polyatomic ion equals the charge on the ion.
 - N = +5 and O = -2 in NO₃⁻, (+5) + 3(-2) = -1

- Oxidation state of oxygen is -2, except in the O₂²⁻ ion, it is -1
- Oxidation state of hydrogen is +1, except in metal hydrides it is -1.
- Group I metals have an oxidation state of +1 in all their compounds.
 - Na = +1 in NaCl
- Group II metals have an oxidation state of +2 in all their compounds.
 - Mg = +2 in MgCl₂
- In their compounds, nonmetals have oxidation states according to the table on the right.

Nonmetal	Oxidation State	Example
Fluorine	-1	MgF ₂ <small>-1 ox state</small>
Hydrogen	+1	H ₂ O <small>+1 ox state</small>
Oxygen	-2	CO ₂ <small>-2 ox state</small>
Group 7A	-1	CCl ₄ <small>-1 ox state</small>
Group 6A	-2	H ₂ S <small>-2 ox state</small>
Group 5A	-3	NH ₃ <small>-3 ox state</small>

- The reactant that reduces an element in another reactant is called the **reducing agent**.
 - The reducing agent contains the element that is oxidized.
- The reactant that oxidizes an element in another reactant is called the **oxidizing agent**.
 - The oxidizing agent contains the element that is reduced.

Redox Reactions

- All the reactions which involve a change in oxidation number of the species involved are redox reactions.
- Double replacement reactions are not redox reactions.

Decomposition reactions

- These reactions are just the opposite of combination reactions.
- a. Hydroxide decompose to give metal oxide and water:

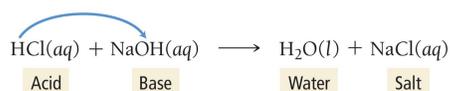
$$\text{Mg}(\text{OH})_2(\text{aq}) \rightarrow \text{MgO}(\text{s}) + \text{H}_2\text{O}(\text{l})$$
- b. Metal carbonates decompose into metal oxide and carbon dioxide

$$\text{MgCO}_3(\text{s}) \rightarrow \text{MgO}(\text{s}) + \text{CO}_2(\text{g})$$
- c. All metal chlorates when heated decompose into metal chlorides and oxygen gas :

$$2\text{KClO}_3(\text{s}) \rightarrow 2\text{KCl}(\text{s}) + 3\text{O}_2(\text{g})$$

Acid–Base Reactions

- When an acid and a base react, the reaction is called a **neutralization reaction**.
- Moles of acid = Moles of base



Combination Reactions

- Metals react with oxygen to give metallic oxides:

$$\text{Mg}(\text{s}) + \text{O}_2(\text{g}) \rightarrow \text{MgO}(\text{s})$$
- Metallic oxides react with water to give hydroxides:

$$\text{MgO}(\text{s}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{Mg}(\text{OH})_2$$
- Metallic oxides react with carbon dioxide to give carbonates:

$$\text{MgO}(\text{s}) + \text{CO}_2(\text{g}) \rightarrow \text{MgCO}_3(\text{s})$$

Combustion reactions

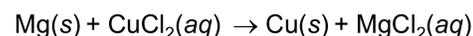
– Common example, hydrocarbon fuel reacts with oxygen to produce carbon dioxide and water

– Consider:



Single replacement reactions

- A common reaction: active metal replaces (displaces) a metal ion from a solution.



- The activity series of metals is useful in order to predict the outcome of the reaction.

TABLE 4.6 Activity Series

Element	Oxidation Half-Reaction
Lithium	Li → Li ⁺ + e ⁻
Potassium	K → K ⁺ + e ⁻
Barium	Ba → Ba ²⁺ + 2e ⁻
Calcium	Ca → Ca ²⁺ + 2e ⁻
Sodium	Na → Na ⁺ + e ⁻
Magnesium	Mg → Mg ²⁺ + 2e ⁻
Aluminum	Al → Al ³⁺ + 3e ⁻
Manganese	Mn → Mn ²⁺ + 2e ⁻
Zinc	Zn → Zn ²⁺ + 2e ⁻
Chromium	Cr → Cr ³⁺ + 3e ⁻
Iron	Fe → Fe ²⁺ + 2e ⁻
Cadmium	Cd → Cd ²⁺ + 2e ⁻
Cobalt	Co → Co ²⁺ + 2e ⁻
Nickel	Ni → Ni ²⁺ + 2e ⁻
Tin	Sn → Sn ²⁺ + 2e ⁻
Lead	Pb → Pb ²⁺ + 2e ⁻
Hydrogen	H ₂ → 2H ⁺ + 2e ⁻
Copper	Cu → Cu ²⁺ + 2e ⁻
Silver	Ag → Ag ⁺ + e ⁻
Mercury	Hg → Hg ²⁺ + 2e ⁻
Platinum	Pt → Pt ²⁺ + 2e ⁻
Gold	Au → Au ³⁺ + 3e ⁻

Increasing ease of oxidation

8.8 Gas-Evolution Chemical Reactions

- In a **gas-evolution reaction**, a gas forms, resulting in bubbling.

TABLE 8.3 Types of Compounds That Undergo Gas-Evolution Reactions

Reactant Type	Intermediate Product	Gas Evolved	Example
Sulfides	None	H ₂ S	2 HCl(aq) + K ₂ S(aq) → H ₂ S(g) + 2 KCl(aq)
Carbonates and bicarbonates	H ₂ CO ₃	CO ₂	2 HCl(aq) + K ₂ CO ₃ (aq) → H ₂ O(l) + CO ₂ (g) + 2 KCl(aq)
Sulfites and bisulfites	H ₂ SO ₃	SO ₂	2 HCl(aq) + K ₂ SO ₃ (aq) → H ₂ O(l) + SO ₂ (g) + 2 KCl(aq)
Ammonium	NH ₄ OH	NH ₃	NH ₄ Cl(aq) + KOH(aq) → H ₂ O(l) + NH ₃ (g) + KCl(aq)

Chemical Reactions
Chapter 8

Chap 9

Energy: Capacity to Do Work

Kinetic Energy:
Due to Motion

Potential Energy: Due to
Position or Composition

Thermal Energy:
Associated with
Temperature

Chemical Energy:
Associated with Positions
of Electrons and Nuclei

$KE = \frac{1}{2}mv^2$ - Unit is Joules w/ mass is kg & velocity is m/s

- A **calorie (cal)** is the amount of energy needed to raise the temperature of one gram of water by 1 °C.
 - 1 kcal = energy needed to raise 1000 g of water by 1 °C.
- A food Calorie (Cal) is 1000 calories.
 - 1 kcal is equivalent to **ONE** food Calorie.

Energy Units =	Conversion
1 calorie (cal) =	4.184 joules (J)
1 kilocalorie (kcal) =	1000 calories (cal)
1 food calorie (Cal) =	1 kcal or 1000 calories
1 kilowatt-hour (kWh) =	3.60×10^6 joules (J)

The First Law of Thermodynamics

- The first law of thermodynamics is the law of conservation of energy.
 - This means that the total amount of energy in the universe is constant; it is conserved.
 - Therefore, you can never design a system that will continue to produce energy without some source of energy.
- Thermodynamics is the study of energy that is exchanged between the system and the surroundings.
 - Energy can flow from the system to the surroundings.
 - Energy of the system drops; energy of the surroundings increase.
 - Exothermic**
 - Energy can flow into the system from the surroundings.
 - Energy of the system increases; energy of the surroundings decreases.
 - Endothermic**

$$\Delta E_{\text{universe}} = 0 = \Delta E_{\text{system}} + \Delta E_{\text{surroundings}}$$

- Energy is exchanged between the system and surroundings through heat and work.
 - q = heat (thermal) energy
 - w = work energy
 - q and w are NOT state functions; their values depend on the process.

$$\Delta E = q + w$$

Internal Energy

- The **internal energy** is the sum of the kinetic and potential energies of all of the particles that compose the system.
- The change in the internal energy of a system only depends on the amount of energy in the system at the beginning and end.
 - A **state function** is a mathematical function whose result only depends on the initial and final conditions, not on the process used.

$$\Delta E = E_{\text{final}} - E_{\text{initial}}$$

$$\Delta E_{\text{reaction}} = E_{\text{products}} - E_{\text{reactants}}$$
- If the reactants have a higher internal energy than the products:
 - ΔE_{sys} is said to be negative because energy flows **OUT** of the system into the surroundings.
- If the reactants have a lower internal energy than the products:
 - ΔE_{sys} is said to be positive because energy flows **INTO** the system from the surroundings.
- The **internal energy** of the system can be treated in a similar manner.
 - Energy flowing **out** of the system is like a withdrawal and therefore carries a negative sign.
 - Energy flowing **into** the system is like a deposit and carries a positive sign.