

## 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 Cl<sub>2</sub> = 0 in 2 Na(s) + Cl<sub>2</sub>(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 NO<sub>3</sub><sup>-</sup>, (+5) + 3(-2) = -1

- Oxidation state of oxygen is -2, except in the O<sub>2</sub><sup>2-</sup> 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 MgCl<sub>2</sub>
- In their compounds, nonmetals have oxidation states according to the table on the right.

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

- 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

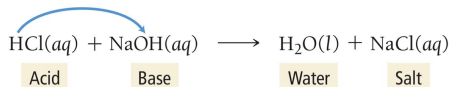
- All the reactions which involve a change in oxidation number of the species involved are redox reactions.
- Double replacement reactions are not redox reactions.

#### 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



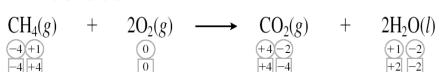
#### Combination Reactions

- 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

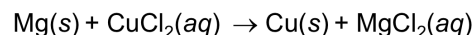
- Common example, hydrocarbon fuel reacts with oxygen to produce carbon dioxide and water

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	$\text{Li} \rightarrow \text{Li}^+ + \text{e}^-$
Potassium	$\text{K} \rightarrow \text{K}^+ + \text{e}^-$
Barium	$\text{Ba} \rightarrow \text{Ba}^{2+} + 2\text{e}^-$
Calcium	$\text{Ca} \rightarrow \text{Ca}^{2+} + 2\text{e}^-$
Sodium	$\text{Na} \rightarrow \text{Na}^+ + \text{e}^-$
Magnesium	$\text{Mg} \rightarrow \text{Mg}^{2+} + 2\text{e}^-$
Aluminum	$\text{Al} \rightarrow \text{Al}^{3+} + 3\text{e}^-$
Manganese	$\text{Mn} \rightarrow \text{Mn}^{2+} + 2\text{e}^-$
Zinc	$\text{Zn} \rightarrow \text{Zn}^{2+} + 2\text{e}^-$
Chromium	$\text{Cr} \rightarrow \text{Cr}^{3+} + 3\text{e}^-$
Iron	$\text{Fe} \rightarrow \text{Fe}^{2+} + 2\text{e}^-$
Cadmium	$\text{Cd} \rightarrow \text{Cd}^{2+} + 2\text{e}^-$
Cobalt	$\text{Co} \rightarrow \text{Co}^{2+} + 2\text{e}^-$
Nickel	$\text{Ni} \rightarrow \text{Ni}^{2+} + 2\text{e}^-$
Tin	$\text{Sn} \rightarrow \text{Sn}^{2+} + 2\text{e}^-$
Lead	$\text{Pb} \rightarrow \text{Pb}^{2+} + 2\text{e}^-$
Hydrogen	$\text{H}_2 \rightarrow 2\text{H}^+ + 2\text{e}^-$
Copper	$\text{Cu} \rightarrow \text{Cu}^{2+} + 2\text{e}^-$
Silver	$\text{Ag} \rightarrow \text{Ag}^+ + \text{e}^-$
Mercury	$\text{Hg} \rightarrow \text{Hg}^{2+} + 2\text{e}^-$
Platinum	$\text{Pt} \rightarrow \text{Pt}^{2+} + 2\text{e}^-$
Gold	$\text{Au} \rightarrow \text{Au}^{3+} + 3\text{e}^-$

## 8.8 Gas-Evolution Chemical Reactions

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

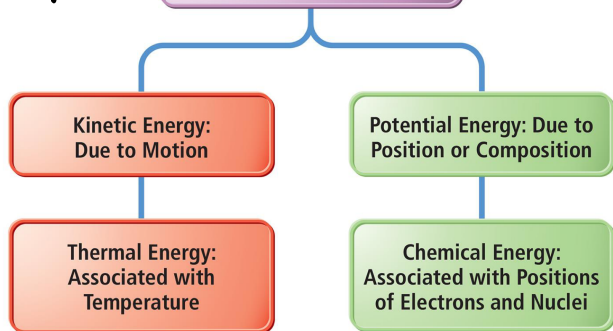
TABLE 8.3 Types of Compounds That Undergo Gas-Evolution Reactions

Reactant Type	Intermediate Product	Gas Evolved	Example
Sulfides	None	H <sub>2</sub> S	$2\text{HCl}(\text{aq}) + \text{K}_2\text{S}(\text{aq}) \rightarrow \text{H}_2\text{S}(\text{g}) + 2\text{KCl}(\text{aq})$
Carbonates and bicarbonates	H <sub>2</sub> CO <sub>3</sub>	CO <sub>2</sub>	$2\text{HCl}(\text{aq}) + \text{K}_2\text{CO}_3(\text{aq}) \rightarrow \text{H}_2\text{O}(\text{l}) + \text{CO}_2(\text{g}) + 2\text{KCl}(\text{aq})$
Sulfites and bisulfites	H <sub>2</sub> SO <sub>3</sub>	SO <sub>2</sub>	$2\text{HCl}(\text{aq}) + \text{K}_2\text{SO}_3(\text{aq}) \rightarrow \text{H}_2\text{O}(\text{l}) + \text{SO}_2(\text{g}) + 2\text{KCl}(\text{aq})$
Ammonium	NH <sub>4</sub> OH	NH <sub>3</sub>	$\text{NH}_4\text{Cl}(\text{aq}) + \text{KOH}(\text{aq}) \rightarrow \text{H}_2\text{O}(\text{l}) + \text{NH}_3(\text{g}) + \text{KCl}(\text{aq})$

Chemical Reactions  
Chapter 8

# Chap 9

Energy: Capacity to Do Work



$KE = \frac{1}{2}mv^2$  - 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) =	$3.60 \times 10^6$ 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**

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

- 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:
  - $\Delta E_{\text{sys}}$  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:
  - $\Delta E_{\text{sys}}$  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.