















Chapter Summary and Review

Key Learning Outcomes



CHAPTER OBJECTIVES	ASSESSMENT
Use the Half-Reaction Method of Balancing Aqueous Redox Equations in Acidic Solution (19.2)	<ul style="list-style-type: none"> Examples 19.1, 19.2 For Practice 19.1, 19.2 Exercises 33, 34, 35, 36
Balance Redox Reactions Occurring in Basic Solution (19.2)	<ul style="list-style-type: none"> Example 19.3 For Practice 19.3 Exercises 37, 38
Calculate Standard Potentials for Electrochemical Cells from Standard Electrode Potentials of the Half-Reactions (19.4)	<ul style="list-style-type: none"> Example 19.4 For Practice 19.4 Exercises 41, 42, 57, 58
Predict Spontaneous Redox Reactions and Sketching Electrochemical Cells (19.4)	<ul style="list-style-type: none"> Example 19.5 For Practice 19.5 Exercises 39, 40, 43, 44, 47, 48, 49, 50
Relate ΔG° and E°_{cell} (19.5)	<ul style="list-style-type: none"> Example 19.6 For Practice 19.6 Exercises 61, 62
Relate E°_{cell} and K (19.5)	<ul style="list-style-type: none"> Example 19.7 For Practice 19.7 Exercises 63, 64, 65, 66, 67, 68
Calculate E_{cell} under Nonstandard Conditions (19.6)	<ul style="list-style-type: none"> Example 19.8 For Practice 19.8 Exercises 69, 70, 71, 72, 73, 74
Predict the Products of Electrolysis Reactions (19.8)	<ul style="list-style-type: none"> Example 19.9 For Practice 19.9 Exercises 87, 88, 89, 90, 91, 92
Use Stoichiometry to Calculate the Quantity of Reactants Consumed or Products Produced in an Electrolytic Cell (19.8)	<ul style="list-style-type: none"> Example 19.10 For Practice 19.10 Exercises 95, 96, 97, 98

Key Terms

Section 19.3

electrical current 
electrochemical cell 
voltaic (galvanic) cell 
electrolytic cell 
half-cell 
electrode
ampere (A) 
potential difference 
volt (V) 
electromotive force (emf) 
cell potential (cell emf) (E_{cell}) 
standard cell potential (standard emf) (E_{cell}°) 
anode 
cathode 
salt bridge 

Section 19.4

standard electrode potential 
standard hydrogen electrode (SHE) 








Section 19.5

Faraday's constant (F) 

Section 19.6

Nernst equation 

Section 19.7

dry-cell battery 
alkaline battery 
lead–acid storage battery 
nickel–cadmium (NiCad) battery 
nickel–metal hydride (NiMH) battery 
lithium ion battery 
fuel cell 

Section 19.8

electrolysis 

Section 19.9

corrosion 

Key Concepts

Lightning and Batteries (19.1)

- Lightning is the massive flow of electrons from the base of a thundercloud (which is negatively charged) to the ground (which is positively charged).
- In a battery, electrons flow from the negatively charged end to the positively charged end.

Balancing Oxidation–Reduction Equations (19.2)

- Oxidation is the loss of electrons and corresponds to an increase in oxidation state; reduction is the gain of electrons and corresponds to a decrease in oxidation state.
- We can balance redox reactions using the half-reaction method, in which the oxidation and reduction reactions are balanced separately and then added. This method differs slightly for redox reactions in acidic and in basic solutions.

Voltaic (or Galvanic) Cells: Generating Electricity from Spontaneous Chemical Reactions (19.3)

- A voltaic electrochemical cell separates the reactants of a spontaneous redox reaction into two half-cells that are connected by a wire and a means to exchange ions so that electricity is generated.
- In an electrochemical cell, the electrode where oxidation occurs is the anode and the electrode where reduction occurs is the cathode; electrons flow from the anode to the cathode.
- The rate of electrons flowing through a wire is measured in amperes (A), and the cell potential is measured in volts (V).
- A salt bridge allows ions to flow between the half-cell solutions and prevents the buildup of charge.
- Cell diagram or line notation is a technique for symbolizing electrochemical cells concisely by separating the components of the reaction using lines or commas.

Standard Electrode Potentials (19.4)

- The electrode potentials of half-cells are measured in relation to that of a standard hydrogen electrode, which is assigned an electrode potential of zero under standard conditions (solution concentrations of 1 M, gas pressures of 1 atm, and 25 °C).
- A species with a highly positive E° has a strong tendency to attract electrons and undergo reduction (and is therefore an excellent oxidizing agent).
- A species with a highly negative E° has a strong tendency to repel electrons and undergo oxidation (and is therefore an excellent reducing agent).

Cell Potential, Free Energy, and the Equilibrium Constant (19.5)

- In a spontaneous reaction, E°_{cell} is positive, the change in free energy (ΔG°) is negative, and the equilibrium constant (K) is greater than 1.
- In a nonspontaneous reaction, E°_{cell} is negative, ΔG° is positive, and K is less than 1.
- Because E°_{cell} , ΔG° , and K all relate to spontaneity, we can derive equations relating all three quantities.

Cell Potential and Concentration (19.6)

- The standard cell potential (E°_{cell}) is related to the cell potential (E_{cell}) by the Nernst equation, $E_{\text{cell}} = E^\circ_{\text{cell}} - (0.0592 \text{ V/n}) \log \mathcal{Q}$.
- As shown by the Nernst equation, E_{cell} is related to the reaction quotient (\mathcal{Q}); E_{cell} equals zero when \mathcal{Q} equals K .
- In a concentration cell, the reactions at both electrodes are identical and electrons flow because of a difference in concentration.

Batteries: Using Chemistry to Generate Electricity (19.7)

- Batteries are packaged voltaic cells.
- Dry-cell batteries, including alkaline batteries, do not contain large amounts of water.
- The reactions in rechargeable batteries, such as lead–acid storage, nickel–cadmium, nickel–metal hydride, and lithium ion batteries, can be reversed.
- Fuel cells are similar to batteries except that fuel-cell reactants must be continually replenished from an external source.

Electrolysis: Driving Nonspontaneous Chemical Reactions with Electricity (19.8)

- An electrolytic electrochemical cell differs from a voltaic cell in that (1) an electrical charge is used to drive

the reaction, and (2) although the anode is still the site of oxidation and the cathode the site of reduction, they are represented with signs opposite those of a voltaic cell (anode +, cathode -).

- In electrolysis reactions, the anion is oxidized; if there is more than one anion, the anion with the more negative E° is oxidized.
- We can use stoichiometry to calculate the quantity of reactants consumed or products produced in an electrolytic cell.

Corrosion: Undesirable Redox Reactions (19.9)

- Corrosion is the undesired oxidation of metal by environmental oxidizing agents.
- When some metals such as aluminum oxidize, they form a stable compound that prevents further oxidation. Iron, however, does not form a structurally stable compound when oxidized and therefore rust flakes off and exposes more iron to corrosion.
- Iron corrosion can be prevented by preventing water contact, minimizing the presence of electrolytes and acids, or using a sacrificial electrode.

Key Equations and Relationships

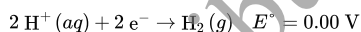
Definition of an Ampere (19.3)

$$1 \text{ A} = 1 \text{ C/s}$$

Definition of a Volt (19.3)

$$1 \text{ V} = 1 \text{ J/C}$$

Standard Hydrogen Electrode (19.4)



Equation for Cell Potential (19.4)

$$E_{\text{cell}}^\circ = E_{\text{cathode}}^\circ - E_{\text{anode}}^\circ$$

Relating ΔG° and E_{cell}° (19.5)

$$\Delta G^\circ = -nFE_{\text{cell}}^\circ \quad F = \frac{96,485 \text{ C}}{\text{mol e}^-}$$

Relating E_{cell}° and K (19.5)

$$E^\circ = \frac{0.0592 \text{ V}}{n} \log K \quad (\text{at } 25^\circ \text{C})$$

The Nernst Equation (19.6)

$$E_{\text{cell}} = E_{\text{cell}}^\circ - \frac{0.0592 \text{ V}}{n} \log Q \quad (\text{at } 25^\circ \text{C})$$

Not for Distribution