Chapter 17
Oxidation-Reduction
Advanced Chemistry

17.1 Oxidation Number

Learning Objective	

Assign oxidation numbers to the atoms in a compound.	• Rules for Assigning Oxidation Numbers
	  ○ All elements in their free state (uncombined with other elements) have an oxidation number of zero (e.g., Na, Cu, Mg, H_2, O_2, Cl_2, N_2).
	  ○ H is +1, except in metal hydrides, where it is -1 (e.g., NaH, CaH_2).
	  ○ O is -2, except in peroxides, where it is -1, and in OF_2, where it is +2.
	  ○ The metallic element in an ionic compound has a positive oxidation number.
	  ○ In covalent compounds the negative oxidation number is assigned to the most electronegative atom.
	  ○ The algebraic sum of the oxidation numbers of the elements in a compound is zero.
	  ○ The algebraic sum of the oxidation number of the elements in a polyatomic ion is equal to the charge of the ion.
	• PROBLEM-SOLVING STRATEGY: Finding the Oxidation Number of an Element in a Compound
	  1. Write the oxidation number of each known atom above and below the atom in the formula.
	  H_2SO_4 (Example 17.2 pg. 392)
	$$\begin{matrix} +1 & & -2 \\ H_2 & S & O_4 \\ +1 & & -2 \end{matrix}$$
	  2. Multiply each oxidation number by the number of atoms of that element in the compound.
	$$\begin{matrix} +1 & & -2 \\ H_2 & S & O_4 \\ 2(+1) & & 4(-2) \end{matrix}$$
	  3. Write an expression indicating the sum of all the oxidation numbers in the compound. Remember: The sum of the oxidation numbers in a compound must equal zero, and in a polyatomic ion the ion charge.
	$$\begin{matrix} +1 & & -2 \\ H_2 & S & O_4 \\ 2(+1)+ & 1(x)+ & 4(-2)=0 \end{matrix}$$
	  x = +6, therefore
	$$\begin{matrix} +1 & +6 & -2 \\ H_2 & S & O_4 \\ 2(+1)+ & 1(x)+ & 4(-2)=0 \end{matrix}$$
	  the oxidation of H = +1, S = +6, and O = -2

Key Terms	
Oxidation number	A small number representing the state of oxidation of an atom. • For an ion, it is the positive or negative charge on the ion; • for covalently bonded atoms, it is a positive or negative number assigned to the more electronegative atom; • in free elements, it is zero.
Oxidation state	Also known as oxidation number
Oxidation-reduction	A chemical reaction wherein electrons are transferred form on element to another; also known as redox.
Redox	Also known as oxidation-reduction
Oxidation	An increase in the oxidation number of an atom as a result of losing electrons.
Reduction	A decrease in the oxidation number of an element as a result of gaining electrons.
Oxidizing agent	A substance that causes an increase in the oxidation state of another substance; the oxidizing agent is reduced during the course of the reaction.
Reducing agent	A substance that causes a decrease in the oxidation state of another substance; the reducing agent is oxidized during the course of a reaction.

17.2 Balancing Oxidation-Reduction Equations

Learning Objective	
Balance equations for oxidation-reduction reactions.	• PROBLEM-SOLVING STRATEGY: Using Charge in Oxidation Numbers to Balance Oxidation-Reduction Reactions 1. Assign oxidation numbers for each element to identify the elements being oxidized and reduced. Look for those elements that have changed oxidation number. 2. Write two half-reactions using only the elements that have changed oxidation numbers. One half-reaction must produce electrons, and the other must use electrons. 3. Multiply the half-reactions by the smallest whole numbers that will make the electrons lost by oxidation equal the electrons gained by reduction. 4. Transfer the coefficient in front of each substance in the balance half-reactions to the corresponding substance in the original equation. 5. Balance the remaining elements that are not oxidized or reduced to give the final balanced equation. 6. Check to make sure both sides of the equation have the same number of atoms of each element.

17.3 Balancing Ionic Redox Equations

Learning Objective	
Balance equations for ionic oxidation-reduction reactions.	• PROBLEM-SOLVING STRATEGY: Ion-Electron Strategy for Balancing Oxidation-Reduction Reactions 1. Write the two half-reactions that contain the elements being oxidized and reduced using the entire formula of the ion or molecule. 2. Balance the elements other than oxygen and hydrogen. 3. Balance oxygen and hydrogen. **Acidic solution**: For reactions in acidic solution, use H^+ and H_2O to balance oxygen and hydrogen. For each oxygen needed, use one H_2O. Then add H^+ as needed to balance the hydrogen atoms. Basic solution: For reactions in alkaline solutions, first balance as though the reaction were in an acidic solution, using Steps 1-3. Then add as many OH^- ions to each side of the equation as there are H^+ ions in the equation. Now combine the H^+ and OH^- ions into water (e.g., 4 H^+ and 4 OH^- give 4 H_2O). Rewrite the equation, canceling equal numbers of water molecules that appear on opposite sides of the equation. 4. Add electrons (e-) to each half-reaction to bring them into electrical balance. 5. Since the loss and gain or electrons must be equal, multiply each half-reaction by the appropriate number to make the number of electrons the same in each half-reaction. 6. Add the two half-reactions together, canceling electrons and any other identical substances that appear on opposite sides of the equation.

17.4 Activity Series of Metals

Learning Objective	
Use the activity series of metals to predict whether a reaction will occur.	• PROBLEM-SOLVING STRATEGY: Using the Activity Series to Predict Reactions o The reactivity of the metals listed decreases from top to bottom. o A free metal can displace the ion of a second metal form solution, provided that the free metal is above the second metal in the activity series. o Free metals above hydrogen react with nonoxidizing acids in solution to liberate hydrogen gas. o Free metals below hydrogen do not liberate hydrogen from acids. o Conditions such as temperature and concentration may affect the relative position of some of these elements.
Key Term	
Activity series of metals	A listing of metallic elements in descending order of reactivity.

17.5 Electrolytic and Voltaic Cells

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<td>Compare the reactions and functions of electrolytic and voltaic cells.</td>
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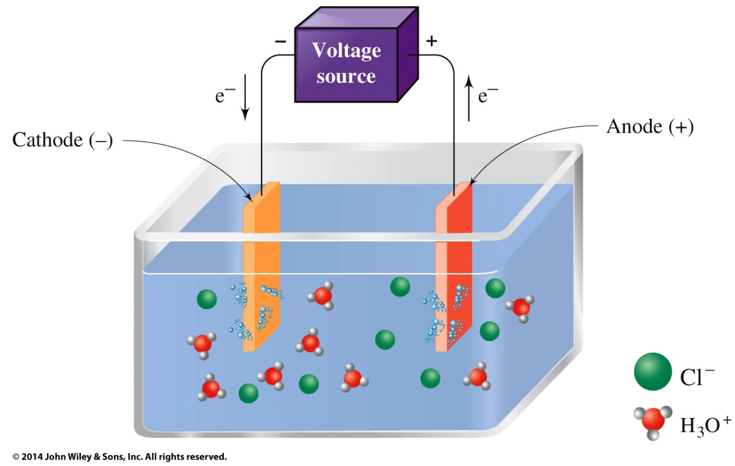
- Electrolytic cells use electrical energy to produce a chemical reaction.



The cell consists of a source of direct current (a battery) connected to two electrodes that are immersed in a solution of hydrochloric acid. The negative electrode is called the cathode because cations are attracted to it. The positive electrode is called the anode because anions are attracted to it. The cathode is attached to the negative pole and the anode to the positive pole of the battery. The battery supplies electrons to the cathode.

When the electric circuit is completed, positive hydronium ions (H_3O^+) migrate to the cathode, where they pick up electrons and evolve as hydrogen gas. At the same time the negative chloride ions (Cl^-) migrate to the anode, where they lose electrons and evolve as chlorine gas.

Reaction at the cathode:

$H_3O^+ + 1\ e^- \rightarrow H^0 + H_2O$ (reduction)

$H^0 + H^0 \rightarrow H_2$

Reaction at the anode:

$Cl^- \rightarrow Cl^0 + 1\ e^-$ (oxidation)

$Cl^0 + Cl^0 \rightarrow Cl_2$

$2\ HCl(aq) \xrightarrow{electrolysis} H_2(g) + Cl_2(g)$ (net reaction)

NOTE: oxidation always occurs at the anode and reduction at the cathode

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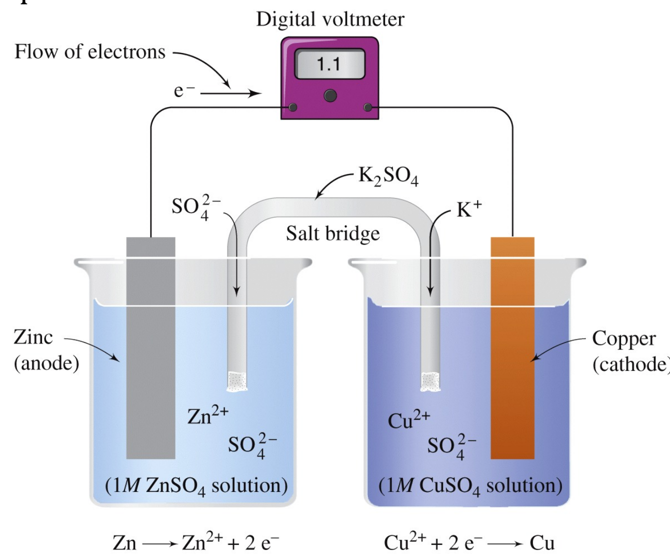
- Voltaic cells (also known as galvanic cells) produce electric current form a spontaneous chemical reaction



The driving force responsible for the electric current in the zinc-copper cell originates in the great tendency of zinc atoms to lose electrons relative to the tendency of copper(II) ions to gain electrons.

In the cell above, zinc atoms lose electrons and are converted to zinc ions at the zinc electrode surface; the electrons flow through the wire (external circuit) to the copper electrode. Here copper(II) ions pick up electrons and are reduced to copper atoms, which plate out on the copper electrode. Sulfate ions flow from the CuSO4 solution via the salt bridge into the ZnSO4 solution (internal circuit) to complete the circuit.

Reaction at the anode:

$Zn^0(s) \rightarrow Zn^{2+}(aq) + 2\ e^-$ (oxidation)

Reaction at the cathode:

$Cu^{2+}(aq) + 2\ e^- \rightarrow Cu^0(s)$ (reduction)

$Zn^0(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu^0(s)$ (net ionic reaction)

$Zn(s) + CuSO_4(aq) \rightarrow ZnSO_4(aq) + Cu(s)$ (overall reaction)

Key Terms	
Electrolysis	The process whereby electrical energy is used to bring about a chemical change.
Electrolytic cell	An electrolysis apparatus in which electrical energy from an outside source is used to produce a chemical change.
Cathode	The electrode where reduction occurs in an electrochemical reaction.
Anode	The electrode where oxidation occurs in an electrochemical reaction.
Voltaic cell	A cell that produces electric current from a spontaneous chemical reaction.