Pearson

Chapter Summary and Review

Key Learning Outcomes

CHAPTER OBJECTIVES	ASSESSMENT
Write Electron Configurations (3.3)	• Example 3.1 For Practice 3.1 Exercises 45, 46, 49, 50
Write Orbital Diagrams (3.3)	• Example 3.2 For Practice 3.2 Exercises 47, 48
Differentiate Between Valence Electrons and Core Electrons (3.4)	• Example 3.3 For Practice 3.3 Exercises 55, 56, 57, 58, 59, 60
Write Electron Configurations from the Periodic Table (3.4)	• Example 3.4 For Practice 3.4 For More Practice 3.4 Exercises 49, 50, 51, 52
Predict the Charge of Ions (3.5)	• Example 3.5 For Practice 3.5 Exercises 63, 64
Use Periodic Trends to Predict Atomic Size (3.6)	• Example 3.6 For Practice 3.6 For More Practice 3.6 Exercises 71, 72, 73, 74
Write Electron Configurations for Ions (3.7)	• Example 3.7 For Practice 3.7 Exercises 63, 64, 75, 76, 77, 78
Apply Periodic Trends to Predict Ion Size (3.7)	• Example 3.8 For Practice 3.8 For More Practice 3.8 Exercises 79, 80, 81, 82
Apply Periodic Trends to Predict Relative Ionization Energies (3.7)	• Example 3.9 For Practice 3.9 For More Practice 3.9 Exercises 83, 84, 85, 86, 87, 88
Predict Metallic Character Based on Periodic Trends (3.8)	• Example 3.10 For Practice 3.10 For More Practice 3.10 Exercises 89, 90, 91, 92, 93, 94

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Key Terms

Section 3.1

periodic property

Section 3.2

periodic law
main-group element
transition element (or transition metal)
family (or group)

Section 3.3

electron configuration
ground state
orbital diagram
Pauli exclusion principle
degenerate
Coulomb's law
shielding
effective nuclear charge (Z_{eff})
penetration
aufbau principle
Hund's rule

Section 3.4

valence electrons
core electrons

Section 3.5

noble gase
metal
nonmetal
metalloid
semiconductor
alkali metal
alkaline earth metal
halogen

Section 3.6

van der Waals radius (nonbonding atomic radius)
covalent radius (bonding atomic radius)
atomic radius

Section 3.7

paramagnetic
diamagnetic
ionization energy (IE)

Section 3.8

electron affinity (EA)

Key Concepts

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Periodic Properties and the Periodic Table (3.1, 3.2)

- The periodic table was developed primarily by Dmitri Mendeleev in the nineteenth century. Mendeleev arranged the elements in a table so that their atomic masses increased from left to right in a row and elements with similar properties fell in the same columns.
- Periodic properties are predictable based on an element's position within the periodic table. Periodic properties include atomic and ionic radius, ionization energy, electron affinity, density, and metallic character.
- Quantum mechanics explains the periodic table by showing how electrons fill the quantum-mechanical orbitals within the atoms that compose the elements.

Electron Configurations (3.3)

- An electron configuration for an atom shows which quantum-mechanical orbitals the atom's electrons occupy. For example, the electron configuration of helium $(1s^2)$ indicates that helium's two electrons exist within the $1s$ orbital.
- The order of filling quantum-mechanical orbitals in multi-electron atoms is: $1s$ $2s$ $2p$ $3s$ $3p$ $4s$ $3d$ $4p$ $5s$ $4d$ $5p$ $6s$.
- According to the Pauli exclusion principle, each orbital can hold a maximum of two electrons (with opposing spins).
- According to Hund's rule, orbitals of the same energy first fill singly with electrons with parallel spins before pairing.

Electron Configurations and the Periodic Table (3.4)

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- An atom's outermost electrons (valence electrons) are most important in determining the atom's properties.
- Because quantum-mechanical orbitals fill sequentially with increasing atomic number, we can predict the electron configuration of an element from its position in the periodic table.

Electron Configurations and the Properties of Elements (3.5)

- The most stable (or chemically unreactive) elements in the periodic table are the noble gases. These elements have completely full principal energy levels, which have particularly low potential energy compared to other possible electron configurations.
- Elements on the left side and in the center of the periodic table are metals and tend to lose electrons when they undergo chemical changes.
- Elements on the upper right side of the periodic table are nonmetals and tend to gain electrons when they undergo chemical changes.
- Elements with one or two valence electrons are among the most active metals, readily losing their valence electrons to attain noble gas configurations.
- Elements with six or seven valence electrons are among the most active nonmetals, readily gaining enough electrons to attain a noble gas configuration.
- Many main-group elements form ions with noble gas electron configurations.

Effective Nuclear Charge and Periodic Trends in Atomic Size (3.6)

- The size of an atom is largely determined by its outermost electrons. As we move down a column in the periodic table, the principal quantum number (n) of the outermost electrons increases, resulting in successively larger orbitals and therefore larger atomic radii.
- As we move across a row in the periodic table, atomic radii decrease because the effective nuclear charge—the net or average charge experienced by the atom's outermost electrons—increases.
- The atomic radii of the transition elements stay roughly constant as we move across each row because electrons are added to the $n_{\text{highest}} - 1$ orbitals, while the number of highest n electrons stays roughly constant.

Ion Properties (3.7)

- We determine the electron configuration of an ion by adding or subtracting the corresponding number of electrons to the electron configuration of the neutral atom.
- For main-group ions, the order of removing electrons is the same as the order in which they are added in building up the electron configuration.
- For transition metal atoms, *ns* electrons are removed before $(n-1)\,d$ electrons.
- The radius of a cation is much *smaller* than that of the corresponding atom, and the radius of an anion is much *larger* than that of the corresponding atom.
- The first ionization energy—the energy required to remove the first electron from an atom in the gaseous state—generally decreases as we move down a column in the periodic table and increases when we move to the right across a row.
- Successive ionization energies increase smoothly from one valence electron to the next, but the ionization energy increases dramatically for the first core electron.

Electron Affinities and Metallic Character (3.8)

- Electron affinity—the energy associated with an element in its gaseous state gaining an electron—does not show a general trend as we move down a column in the periodic table, but it generally becomes more negative (more exothermic) to the right across a row.
- Metallic character—the tendency to lose electrons in a chemical reaction—generally increases down a column in the periodic table and decreases to the right across a row.

Key Equations and Relationships

Order of Filling Quantum-Mechanical Orbitals (3.3)

$$1s\ 2s\ 2p\ 3s\ 3p\ 4s\ 3d\ 4p\ 5s\ 4d\ 5p\ 6s$$

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