General Organic and Biochemistry 10th edition Denniston Solutions Manual

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Chapter 2
The Structure of the Atom and the Periodic Table
Solutions to the Even-Numbered Questions and Problems

In-Chapter Questions and Problems

2.2
a. Phosphorus, atomic number = 15, therefore 15 protons and 15 electrons.
Mass number - atomic number = 30 - 15 = 15, therefore 15 neutrons.
b. Sulfur, atomic number = 16, therefore 16 protons and 16 electrons.
Mass number - atomic number = 32 - 16 = 16, therefore 16 neutrons.
c. Chlorine, atomic number = 17, therefore 17 protons and 17 electrons.
Mass number - atomic number = 35 - 17 = 18, therefore 18 neutrons.

2.4 Bohr described an atomic orbital as a circular path followed by the electron. An orbital is a region of space in the atom, described by a mathematical function, which predicts the probability of finding an electron in the atom.

2.6
a. S, nonmetal
b. O, nonmetal
c. P, nonmetal
d. N, nonmetal

2.8
a. Ar (argon)
b. C (carbon)
c. B (boron)
d. Ar (argon)

2.10
a. Magnesium, atomic number = 12, mass = 24.31 amu
b. Neon, atomic number = 10, mass = 20.18 amu
c. Selenium, atomic number = 34, mass 78.96 amu

2.12
a. Nickel (Ni) has the atomic number 28. Therefore, Ni^{2+} has 28 protons. A neutral nickel atom has 28 electrons and must lose 2 electrons to form the Ni^{2+} cation. As a result, Ni^{2+} has 26 electrons.
b. Bromine (Br) has the atomic number 35. Therefore, Br^{-} has 35 protons. A neutral bromine atom has 35 electrons and must gain 1 more electrons to form the Br^{-} anion. As a result, Br^{-} has 36 electrons.
c. Nitrogen (N) has the atomic number 7. Therefore, N^{3-} has 7 protons. A neutral nitrogen atom has 7 electrons and must gain 3 more electrons to form the N^{3-} anion. As a result, N^{3-} has 10 electrons.

2.14
a. Cl^{-} has 18 electrons, Ar has 18 electrons; isoelectronic
b. Na^{+} has 10 electrons, Ne has 10 electrons; isoelectronic
c. Mg^{2+} has 10 electrons, Na^{+} has 10 electrons; isoelectronic
d. Li^{+} has 2 electrons, Ne has 10 electrons; not isoelectronic
e. O^{2-} has 10 electrons, F^{-} has 10 electrons; isoelectronic
f. N^{3-} has 10 electrons, Cl- has 18 electrons; not isoelectonic

2.16 a. Rb^+ has 36 electrons. $1s^22s^22p^63s^23p^64s^23d^{10}4p^6$ [Kr]
b. Sr^{2+} has 36 electrons. $1s^22s^22p^63s^23p^64s^23d^{10}4p^6$ [Kr]
c. S^{2-} has 18 electrons. $1s^22s^22p^63s^23p^6$ [Ar]
d. I^- has 54 electrons $1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^24d^{10}5p^6$ [Xe]

2.18 a. (smallest) F, Br, I (largest)
b. (lowest) I, Br, F (highest)
c. (lowest) I, Br, F (highest)

End-of-Chapter Questions and Problems

2.20 The mass of the electrons is too small to be significant in comparison to that of the nucleus.

2.22 The isotopic symbol ${}^{13}_{7}C$ is incorrect. The atomic number for carbon is 6. Nitrogen is the element with the atomic number 7.

2.24 a. False. Atoms with a different number of protons are different elements.
b. True
c. False. The mass number of an atom is due to the mass of its protons and neutrons.

2.26 Atoms A and C are isotopes because they both contain the same number of protons but different numbers of neutrons.

2.28 a. atomic number = 17, hence 17 protons and 17 electrons; 37 –17 = 20 neutrons
b. atomic number = 11, hence 11 protons and 11 electrons; 23 –11 = 12 neutrons
c. atomic number = 36, hence 36 protons and 36 electrons; 84 –36 = 48 neutrons

2.30 Atomic number = number of protons = 19, therefore the atom is K (potassium).
Mass number = number of protons + number of neutrons = 19 + 20 = 39

Therefore, the symbol is ${}^{39}_{19}K$

2.32 a. In-115 has an atomic number of 49, so In-115 has 49 protons.
b. In-115 has 115-49 = 66 neutrons.

2.34 a. Iodine has an atomic number of 53, so iodine has 53 protons.
b. Iodine – 131 has 131 – 53 = 78 neutrons.

2.36 a. 1 proton: atomic number = 1, so the element is H
2 neutrons: mass number = 1 + 2 = 3
atomic symbol: ${}^{3}_{1}H$

b. 92 protons: atomic number = 92, so the element is U
146 neutrons: mass number = 92 + 146 = 238
atomic symbol: $^{238}_{92}U$

2.38 Step 1. Convert each percentage to a decimal fraction.

$$7.49\% \text{ Lithium-6} \times \frac{1}{100\%} = 0.0749 \text{ Lithium-6}$$

$$92.51\% \text{ Lithium-7} \times \frac{1}{100\%} = 0.9251 \text{ Lithium-7}$$

Step 2. Multiply the decimal fraction of each isotope by the mass of that isotope to determine the isotopic contribution to the atomic mass.

Contribution to atomic mass by lithium-6	=	fraction of all Li atoms that are lithium-6	X	mass of a lithium-6 atom
	=	0.0749	X	6.0151 amu
	=	0.4505 amu		
Contribution to atomic mass by lithium-7	=	fraction of all Li atoms that are lithium-7	X	mass of a lithium-7 atom
	=	0.9251	X	7.0160 amu
	=	6.4905 amu		

Step 3. The weighted average is:

Atomic mass of naturally occurring Li	=	contribution of lithium-6	+	contribution of lithium-7
	=	0.4505 amu	+	6.4905 amu
	=	6.9410 amu		
	=	6.94 amu (three significant figures)		

2.40 Points of Dalton's theory that are no longer current:

- An atom cannot be created, divided, destroyed, or converted to any other type of atom. (This point was disproved by the discovery of nuclear fusion, fission, and radioactivity.)
- Atoms of a particular element have identical properties. (This point was disproved by the discovery of isotopes.)

2.42 Crookes used a cathode ray tube. He observed particles emitted by the cathode and traveling toward the anode. This ray was deflected by an electric field. Thomson measured the curvature of the ray influence by the electric and magnetic fields. This measurement provided the mass-to-charge ratio of the negative particle. Thomson also gave the particle the name, electron.

2.44 a. Thomson measured the mass-to-charge ratio of cathode rays and called the negative particle responsible for cathode rays "electrons."

b. Rutherford interpreted the "gold foil experiment" from which he developed the structure of the atom as containing a dense, positive nucleus surrounded by a diffuse sphere of electrons; maintained that the nucleus was responsible for the mass of the atom and the nucleus accounted for only a small fraction of the total volume of the atom.

c. Geiger provided the basic experimental evidence for the existence of the nucleus. A small, dense, positively charged region within the atom was indicated by his alpha-particle scattering experiment.

d. The Bohr theory describes electron arrangement in atoms. Bohr proposed an atomic model that depicted the atom as a nucleus surrounded by fixed energy levels that can be occupied by electrons. He believed that each level was defined by a circular orbit located at some specified distance from the nucleus. Electron promotion from a lower to higher energy level results from absorption of energy that produces an excited state atom. The process of relaxation allows the atom to return to the ground state (the electron falls from a higher to lower energy level) and energy is released.

2.46 Cathode rays are negatively charged. Since opposite charges attract, they deflected toward the positive pole.

2.48 An alpha particle has a positive charge. Since like charges repel, it is deflected away from the positively charged nucleus.

2.50 Rutherford interpreted Geiger's experiment and concluded that the atom is principally empty space. Geiger fired alpha particles at a thin metal foil target and found that most alpha particles passed through the foil instead of being deflected. Rutherford interpreted this to mean that most of the volume of each atom was empty space. If matter was evenly distributed throughout the atom, the matter would have interfered with most of the alpha particles.

2.52 The electromagnetic spectrum is the complete range of electromagnetic waves.

2.54 Electromagnetic radiation, or light, is made up of particles. The energy of each particle is inversely proportional to the wavelength of light.

2.56 The energy of each particle is inversely proportional to the wavelength of light.

2.58 Infrared radiation has a longer wavelength than ultraviolet radiation. The wavelength of ultraviolet radiation is approximately 10^2 nm, and infrared radiation has wavelengths ranging from 10^3 to 10^5nm.

2.60 Only certain wavelengths that are characteristic of the gas are emitted in the emission spectrum. Different gases have different emission spectra composed of different wavelengths of light.

2.62 The promotion of electrons to higher energy levels requires the absorption of energy.

2.64 Reasons why the Bohr theory did not stand the test of time:

- Although the Bohr theory explains the hydrogen spectrum, it provides only a crude approximation of the spectra for more complex atoms.
- We now believe that electrons could exhibit behavior similar to waves.

2.66 Bohr's model was too simplistic; it confined electrons to narrow regions of space. Consequently, it is limited to simple systems, such as hydrogen.

2.68
a. Calcium (Ca) has an atomic number of 20 and an atomic mass of 40.08.
b. Copper (Cu) has an atomic number of 29 and an atomic mass of 63.55.
c. Cobalt (Co) has an atomic number of 27 and an atomic mass of 58.93.
d. Silicon (Si) has an atomic number of 14 and an atomic mass of 28.09.

2.70 Group IIA or (2) is known collectively as the alkaline earth metals and consists of beryllium (Be), magnesium (Mg), calcium (Ca), strontium (Sr), barium (Ba), and radium (Ra).

2.72 Group VIIIA or (18) is known collectively as the noble gases and consists of helium, neon, argon, krypton, xenon, and radon.

2.74
a. The metals are: Ca, K, Cu, Zn
b. The representative metals are: Ca, K
c. The element that is inert is Kr

2.76 According to the graph prepared for Problem 2.65, francium (Fr, atomic number 87) corresponds to a melting point of approximately 26 ºC. The literature value for the melting point of francium is 27 ºC.

2.78 A sublevel is a part of a principal energy level and is designated *s, p, d,* and *f*. Each sublevel may contain one or more orbitals, regions of space containing a maximum of two electrons with their spins paired.

2.80 A 2s orbital is found in the second principal energy level. A 1s orbital is found in the first principal energy level. A 2s orbital is larger than a 1s orbital.

2.82 a. two s electrons; 2 e$^-$/orbital x 1 orbital = 2 e$^-$

b. six p electrons; 2 e$^-$/orbital x 3 orbitals = 6 e$^-$

c. ten d electrons; 2 e$^-$/orbital x 5 orbitals = 10 e$^-$

2.84 Hund's rule: When there is a set of orbitals of equal energy, each orbital becomes half-filled before any become completely filled.

a. Orbital diagram does not violate Hund's rule.

b. Orbital diagram does not violate Hund's rule.

c. Orbital diagram violates Hund's rule. Since the three 2p orbitals are of equal energy, two of the orbitals should be half-filled instead of having one 2p orbital completely filled.

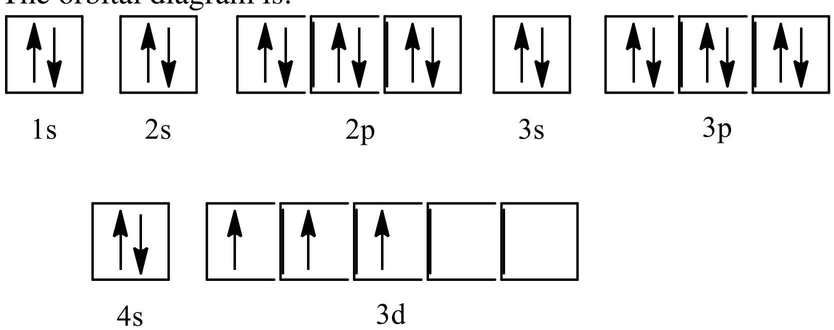
2.86 a. Ca $1s^22s^22p^63s^23p^64s^2$

b. Fe $1s^22s^22p^63s^23p^64s^23d^6$

c. Cl $1s^22s^22p^63s^23p^5$

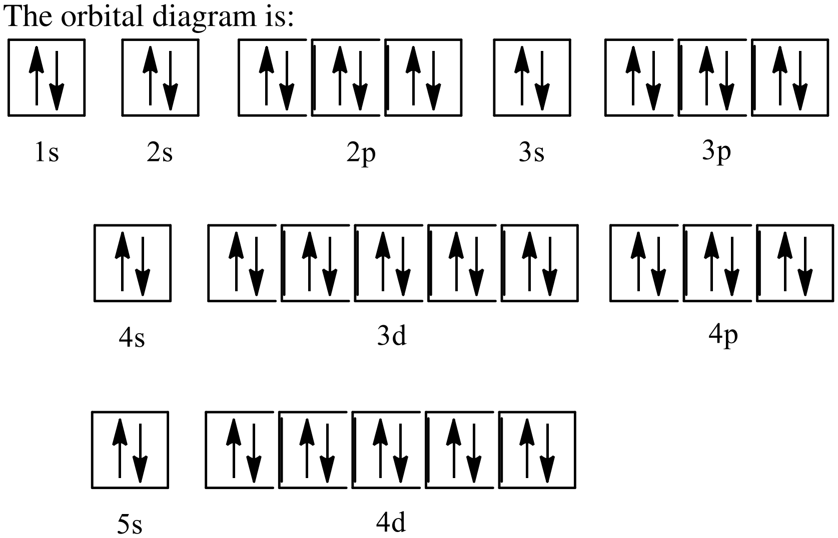
2.88 a. V (23 e$^-$) $1s^22s^22p^63s^23p^64s^23d^3$

The orbital diagram is:



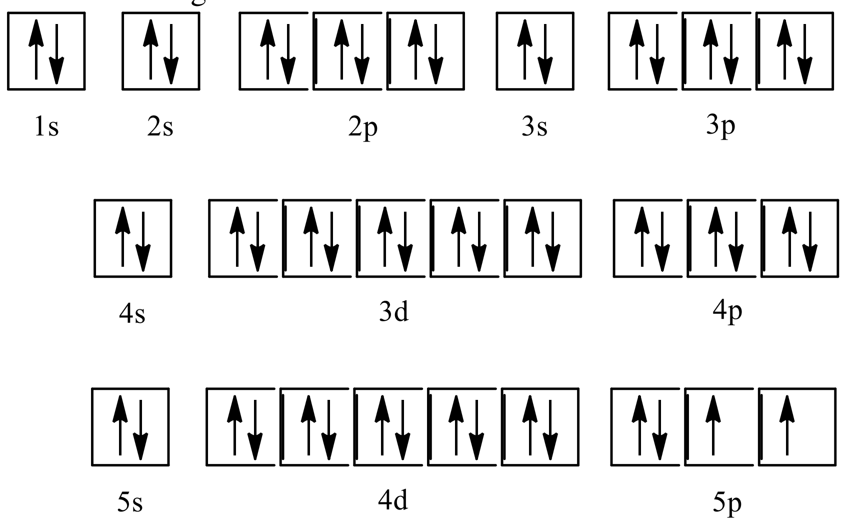
b. Cd (48 e$^-$) $1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^24d^{10}$

The orbital diagram is:



c. Te (52 e⁻) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^4$

The orbital diagram is:



2.90 a., c., and d. are incorrect.
For a., 4 electrons, the element is helium, and the correct electron configuration is:

$$1s^2 2s^2$$

For c., 11 electrons, the element is sodium, and the correct electron configuration is:

$$1s^2 2s^2 2p^6 3s^1$$

For d., 5 electrons, the element is boron, and the correct electron configuration is;

$$1s^2 2s^2 2p^1$$

2.92 The element represented by orbital diagram A is boron. The element represented by orbital diagram B is carbon. The element represented by orbital diagram C is nitrogen.

2.94
a. I has 53 electrons. The noble gas which comes before iodine is Kr. Putting [Kr] in the configuration accounts for the first 36 electrons. The shorthand electron configuration is: $[Kr]5s^2 4d^{10} 5p^5$
b. Al has 13 electrons. The noble gas which comes before aluminum is Ne. Putting [Ne] in the configuration accounts for the first 10 electrons. The shorthand electron configuration is: $[Ne]3s^2 3p^1$
c. V has 23 electrons. The noble gas which comes before vanadium is Ar. Putting [Ar] in the configuration accounts for the first 18 electrons. The shorthand electron configuration is: $[Ar]4s^2 3d^3$

2.96 For representative elements, the number of valence electrons in an atom corresponds to the Roman number of the group in which the atom is found.

2.98 The octet rule can be used to predict the charge of atoms when they become ions. Atoms gain or lose electrons to obtain a noble gas configuration. The ions formed are isoelectronic with the nearest noble gas.

2.100 Nonmetals tend to gain electrons to become negatively charged anions.

2.102 The principal energy level is the same as the period of the periodic table that the element is located in.

	Atom	Total electrons	Valence electrons	Principal energy level number
a.	Mg	12	2	3
b.	K	19	1	4
c.	C	6	4	2
d.	Br	35	7	4
e.	Ar	18	8	3
f.	Xe	54	8	5

2.104 a. Sulfur (S) has the atomic number 16. Therefore, S^{2-} has 16 protons. A neutral sulfur atom has 16 electrons and must gain 2 more electrons to form the S^{2-} anion. As a result, S^{2-} has 18 electrons.
b. Potassium (K) has the atomic number 19. Therefore, K^+ has 19 protons. A neutral potassium atom has 19 electrons and must lose 1 electron to form the K^+ cation. As a result, K^+ has 18 electrons.
b. Cadmium (Cd) has the atomic number 48. Therefore, Cd^{2+} has 48 protons. A neutral cadmium atom has 48 electrons and must lose 2 electrons to form the Cd^{2+} cation. As a result, Cd^{2+} has 46 electrons.

2.106 a. 4 c. 6
b. 5 d. 7

2.108 a. O^{2-} (O gains 2 electrons to attain outermost octet)
b. Br^- (Br gains 1 electron to attain outermost octet)
c. Al^{3+} (Al loses 3 electrons to attain outermost octet)

2.110 a. F^-, $10e^-$; Cl^-, 18 e^-; Not isoelectronic b. K^+, 18 e^-; Ar, $18e^-$; Isoelectronic

2.112 a. Ca^{2+} has 18 electrons $1s^22s^22p^63s^23p^6$ [Ar]
b. Mg^{2+} has 10 electrons $1s^22s^22p^6$ [Ne]
c. K^+ has 18 electrons $1s^22s^22p^63s^23p^6$ [Ar]
d. Cl^- has 18 electrons $1s^22s^22p^63s^23p^6$ [Ar]

2.114 Atomic size increases from top to bottom down a group in the periodic table.

2.116 Electron affinity is the energy change when a single electron is added to a neutral isolated atom.

2.118 Energy is released when an electron is added to an isolated chlorine atom.
$Cl + e^- \rightarrow Cl^- + \text{energy}$

2.120 a. (Smallest) P, Si, Al (Largest)
Atomic size decreases as we go across the periodic table within a period.
b. (Smallest) Al, Ga, In (Largest)
Atomic size increases as we go down a group.
c. (Smallest) Ca, Sr, Ba (Largest)
Atomic size increases as we go down a group.
d. (Smallest) N, P, Sb (Largest)
Atomic size increases as we go down a group.

2.122 a. (Smallest) I, Br, Cl (Largest)
Ionization energy generally decreases as we go down a group.
b. (Smallest) Ra, Mg, Be (Largest)
Ionization energy generally decreases as we go down a group.

2.124 a. (Largest) Cl, P, Mg (Smallest)
Electron affinity generally increases as we go across the periodic table within a period.
b. (Largest) Cl, Br, I (Smallest)
Electron affinity generally decreases as we go down a group

2.126 a. A negative ion is always larger than its parent atom because the positive charge in the nucleus is shared among a larger number of electrons in the ion. Each electron moves further from the nucleus and the volume of the ion increases.
b. A sodium ion has a complete octet of electrons and an electron configuration resembling its nearest noble gas. The ionization energy for sodium is quite low; therefore it is easy for an electron to be removed.

2.128 Ar is larger. Each (Ar and K^+) has the same number of electrons (isoelectronic); however, K^+ has one more positive charge in its nucleus, and this excess positive charge draws all electrons closer to the potassium nucleus, making K^+ smaller than the argon atom.

2 *The Structure of the Atom and the Periodic Table*

Chapter Outline

Learning Goals

1. Describe the properties of protons, neutrons, and electrons.
2. Interpret atomic symbols, and calculate the number of protons, neutrons, and electrons for atoms.
3. Distinguish between the terms atom and isotope, and use isotope notations and natural abundance values to calculate atomic masses.
4. Summarize the history of the development of atomic theory, beginning with Dalton.
5. Describe the role of spectroscopy and the importance of electromagnetic radiation in the development of atomic theory.
6. State the basic postulates of Bohr's theory, its utility, and its limitations.
7. Recognize the important subdivisions of the periodic table: periods, groups (families), metals, and nonmetals.
8. Identify and use the specific information about an element that can be obtained from the periodic table.
9. Describe the relationship between the electronic structure of an element and its position in the periodic table.
10. Write electron configurations, shorthand electron configurations, and orbital diagrams for atoms and ions.
11. Discuss the octet rule, and use it to predict the charges and the numbers of protons and electrons in cations and anions formed from neutral atoms.
12. Utilize the periodic table trends to estimate the relative sizes of atoms and ions, as well as relative magnitudes of ionization energy and electron affinity.

In-Chapter Examples

Example 2.1: ***Determining the Composition of an Atom.*** The atomic symbol for the fluorine atom is used.
Example 2.2: ***Determining Atomic Mass from Two Isotopes.*** Chlorine-35 and chlorine-37 are the isotopes.
Example 2.3: ***Determining Atomic Mass from Three Iostopes.*** Neon-20, neon-21, and neon-22 are the isotopes.
Example 2.4: ***Writing the Electron Configuration.*** The atom is tin.
Example 2.5: ***Writing the Electron Configuration and Orbital Diagram.*** The atom is silicon.
Example 2.6: ***Writing Shorthand Electron Configurations.*** The atom is tin.
Example 2.7: ***Determining Electron Arrangement.*** The atom is silicon.
Example 2.8: ***Determining Proton and Electron Composition of an Ion.*** The ion is Sr^{2+}.
Example 2.9: ***Determining Isoelectronic Ions and Atoms.*** The selenium ion forms.
Example 2.10: ***Writing Shorthand Electron Configurations for Ions.*** The ion is Sb^{3-}.

Chapter Overview

Composition of the Atom

The basic structural unit of an element is the atom, which is the smallest unit of an element that retains the chemical properties of that element. The atom is composed of three primary particles: the electron, the proton, and the neutron.

The atom has two distinct regions. The *nucleus* is a small, dense, positively charged region in the center of the atom composed of positively charged *protons* and uncharged *neutrons*. Surrounding the nucleus is a diffuse region of negative charge occupied by electrons, the source of the negative charge. Electrons are tiny in comparison to protons and neutrons. (**Learning Goal 1**)

The *atomic number* (Z) is equal to the number of protons in the atom. The *mass number* (A) is equal to the sum of the protons and neutrons (the mass of the electrons is insignificant). (**Learning Goal 2**)

Isotopes are atoms of the same element that have different masses because they have different numbers of neutrons (different mass numbers). Aside from potential differences in radioactivity, isotopes have chemical behavior identical to that of any other isotope of the same element. (**Learning Goal 3)**

Development of Atomic Theory

The first experimentally based theory of atomic structure was proposed by John Dalton. Although Dalton pictured atoms as indivisible, the experiments of William Crookes, Eugene Goldstein, and J.J. Thomson indicated that the atom is composed of charged particles: protons and electrons. The third fundamental atomic particle is the neutron. An experiment conducted by Hans Geiger led Ernest Rutherford to propose that the majority of the mass and positive charge of the atom is located in a small, dense region, the *nucleus*, with small, negatively charged electrons occupying a much larger, diffuse space outside of the nucleus. (**Learning Goal 4**)

Light, Atomic Structure, and the Bohr Atom

The study of the interaction of light and matter is termed *spectroscopy*. Light, electromagnetic radiation, travels at a speed of 3.0×10^8 m/s; referred to as the *speed of light*. Light is made up of many wavelengths. Collectively, they comprise the electromagnetic spectrum. Samples of elements emit certain wavelengths of light when an electrical current is passed through the sample. Different elements emit a different pattern (different wavelengths) of light (*emission spectrum*). (**Learning Goal 5**)

Neils Bohr proposed an atomic model that described the atom as a nucleus surrounded by fixed *energy levels* that can be occupied by electrons. He believed that each level was defined by a circular orbit located at a specific distance from the nucleus.

Promotion and relaxation processes are referred to as electronic transitions. Electron promotion resulting from absorption of energy results in an *excited state* atom; the process of relaxation allows the atom to return to the *ground state* by emitting energy in the form of a *photon* (particle of light) with a certain wavelength. (**Learning Goal 6**)

The Periodic Law and the Periodic Table

The periodic law is embodied by Mendeleev's statement, "the elements, if arranged according to their atomic weights (masses), show a distinct *periodicity* (regular variation) of their properties." The periodic table is an organized "map" of the elements that relates their structure to their chemical and physical properties. The periodic table is the result of the periodic law, and provides the basis for prediction of such properties as relative atomic and ionic size, ionization energy, and electron affinity, as well as metallic or non-metallic character and reactivity.

The modern periodic table exists in several forms. The most important variation is in group numbering. The tables in the text use the two most commonly accepted numbering systems.

Periods are the horizontal rows of elements in the periodic table; the columns represent groups or families. The lanthanide series and the actinide series are parts of periods 6 and 7, respectively, and groups that have been named include the *alkali metals*, the *alkaline earth metals*, the *halogens*, and the *noble gases*. Group A elements are called *representative elements*; Group B elements *are transition elements*. Metals, metalloids, and nonmetals can be identified by their location on the periodic table. (**Learning Goal 7)**

The atomic number and the average atomic mass of each element are available from the periodic table. (**Learning Goal 8**)

Electron Arrangement and the Periodic Table

Principle energy levels (n = 1, 2, 3, etc) are regions where electrons may be found. The value that designates the principal energy level is related to the electron's distance from the nucleus. The total electron capacity of a principal level is $2(n)^2$.

The *sublevel* is a set of equal-energy orbitals within a principal energy level. The sublevels are symbolized as *s*, *p*, *d*, and so forth.

An *atomic orbital* is a specific region of a sublevel containing a maximum of two electrons. The electrons in the same orbital must have opposite spins. The number and arrangement of unpaired electrons in an atom are responsible for the magnetic properties of elements.

The electron arrangement (*electron configuration*) and orbital diagrams for an atom can be written using the Aufbau principle, Pauli's exclusion principle and Hund's rule. The shape of the periodic table correlates physical properties of atoms and their electron configurations. Electron configuration of the elements is predictable. Knowing the electron configuration, we can identify valence electrons and begin to predict the kinds of reactions that the elements will undergo. (**Learning Goals 9 and 10**)

Valence Electrons and the Octet Rule

The outermost electrons in an atom are *valence electrons*. For representative elements, the number of valence electrons in an atom corresponds to the group number. Metals tend to have fewer valence electrons than nonmetals.

Elements in the last family, the noble gases, have either two or eight valence electrons. Their most important properties are their extreme stability and lack of reactivity. A full energy level is responsible for this unique stability.

The octet rule tells us that in chemical reactions elements will gain, lose or share the minimum number of electrons necessary to achieve the electron configuration of the nearest inert gas.

Metallic elements tend to form cations and nonmetals form anions that are isoelectronic with their nearest noble gas neighbor. *Ions* are electrically charged particles that result from a gain or loss of one or more electrons by the parent atom. *Anions*, negative ions, are formed by a gain of one or more electrons by the parent atom. *Cations*, positive ions, are formed by a loss of one or more electrons from the parent atom. (**Learning Goal 11**)

Trends in the Periodic Table

Atomic size increases from top to bottom but decreases from left to right in the periodic table. Cations are smaller than the parent atom. Anions are larger than the parent atom. Ions with multiple positive charge are even smaller than their corresponding monopositive ion; ions with multiple negative charge are larger than their corresponding less negative ion.

The energy required to remove an electron from the atom is the ionization energy. Descending a group, the ionization energy decreases. Proceeding across a period, the ionization energy increases.

The energy released when a single electron is added to neutral atom in the gaseous state is known as the electron affinity. Electron affinities generally decrease proceeding down a group and increase proceeding across a period.

Exceptions exist for periodic trends. They are generally small anomalies, and do not detract from the predictive power of the periodic table.
(**Learning Goal 12**)

Hints for Faster Coverage

Although the discussions detailing the historical development of the theory of atomic structure and the importance of ionization energy and electron affinity in ion formation help to give the student a more complete picture, neither is essential for the understanding of the material which is presented in subsequent chapters. Consequently, they may be abbreviated or omitted.

Problems Designed to Support Student Learning

In addition to the "For Further Practice" problems found in each worked example in the text, the "Questions and Problems", located at the end-of-chapter, provide students with a balanced set of problems to test their knowledge of the content of the chapter. The "Multiple Concept Problems", also found at the end-of-chapter, are thought provoking questions designed to challenge students to apply their knowledge to many aspects of a particular topic.

The solutions to the odd-numbered problems are provided in the Student Study Guide and Solution Manual. However, the solutions to the even-numbered problems and the "Multiple Concept Problems" are only available in the Instructor Resource section of the Connect website. Some instructors choose to use the even-numbered problems for assessment. The "Multiple Concept Problems" may also be useful in supporting flipped or team-based learning environments.

It may also help student learning if the instructor emphasizes the importance of active learning strategies such as taking notes by hand, creating flash cards, and developing concept maps.

In-Chapter Perspectives

CHEMISTRY AT THE CRIME SCENE: *Microbial Forensics.* The measurable differences between elemental isotopic oxygen ratios are correlated with their locations. The Federal Bureau of Investigation uses this information to determine where microbes are cultured.

GREEN CHEMISTRY: *Practical Applications of Electromagnetic Radiation*. The practical application of radio waves, X-rays, solar energy, and microwaves is described using a variety of examples.

A HUMAN PERSPECTIVE: *Atomic Spectra and the Fourth of July.* The sights and sounds of a fireworks display result from a chemical reaction that generates the energy necessary to excite a variety of elements resulting in a colorful spectral extravaganza.

A MEDICAL PERSPECTIVE: *Copper Deficiency and Wilson's Disease.* This perspective uses the element copper to describe the biochemical importance of trace minerals in the diet. The role of dietary copper in the function of the central nervous system is discussed, as well as the serious consequences of copper deficiency.

A MEDICAL PERSPECTIVE: *Dietary Calcium.* This perspective describes the essential role that the element calcium plays in our growth and development, as well as our long-term health. It was once believed that our need for calcium diminished significantly once we passed adolescence. This is clearly not true, based on research into the variety of bodily functions in which calcium plays a critical part.

Additional Perspectives

The topic of electromagnetic radiation affords an opportunity to introduce a number of modern applications of light energy. For example, the use of light sensors to regulate blinds and shades in modern office buildings, thus minimizing heating and cooling costs, is an interesting sidelight. The design of lasers, their unique properties, and their applications in laser microsurgery, particularly in retinal repair, could lead to an interesting discussion.

Other perspectives that can be inserted into the lectures dealing with periodic properties might revolve around the special utility of certain elements. For example, silicon and germanium can be discussed in relation to the semiconductor industry; the calculators which students use function because the integrated circuits contain these elements. Also, the unique properties of the transition metals give rise to some useful practical chemistry. For example, metallic mercury, a liquid in the thermometers; thermometers function because of the marked change in volume of mercury with change in temperature. Additional potentially useful topics include copper as a good electrical conductor and many trace metals are essential nutrients.