6

Electronic Structure of Atoms

OVERVIEW OF THE CHAPTER

6.1 Electromagnetic Radiation

Learning Goals: You should be able to:

1. Describe the wave properties and characteristic speed of propagation of radiant energy (electromagnetic radiation).
2. Use the relationship $\lambda v = c$, which relates the wavelength (λ) and the frequency (v) of radiant energy to its speed (c).

6.2 Quantization of Energy

Learning Goals: You should be able to:

1. Explain the essential feature of Planck's quantum theory and use the equation $E = hv$.
2. Explain how Einstein accounted for the photoelectric effect by considering the radiant energy to be a stream of quantized energy packets (photons) striking a metal surface.

6.3 Line Spectra and the Bohr Model

Learning Goals: You should be able to:

1. Explain the origin of the *line spectra*.
2. List the three postulates made by Bohr in his model of the hydrogen atom.
3. Explain the concept of an allowed energy state and how this concept is related to quantum theory.
4. Calculate the energy differences between any two allowed energy states of the electron in a Bohr hydrogen atom.
5. Explain the concept of ionization energy.

6.4, 6.5, 6.6 Principles of Modern Quantum Theory

Learning Goals: You should be able to:

1. Calculate the characteristic wavelength of a particle from a knowledge of its mass and velocity.
2. Describe the uncertainty principle and explain the limitation it places on our ability to define simultaneously the location and momentum of a subatomic particle, particularly an electron.
3. Explain the concepts of orbitals, electron density, probability as used in the quantum-mechanical model of the atom, and the physical significance of Ψ^2.
4. Describe the quantum numbers n, l, and m_l when used to define an orbital and know any limitations placed on their values.
5. Describe the shapes of the s, p, and d orbitals.

6.7 Energies of Orbitals in Many-Electron Atoms

Learning Goals: You should be able to:

1. Explain why electrons with the same value of principal quantum number (n) but with different values of the azimuthal quantum number (l) possess different energies.

6.8 Electronic Structure of Many-Electron Atoms

Learning Goals: You should be able to:

1. Explain the concepts of electron spin and the electron spin quantum number.
2. State the Pauli exclusion principle and Hund's rule, and illustrate how they are used in writing the electronic structures of the elements.
3. Write the electron configuration for any element.
4. Write the orbital diagram representation for electron configurations of atoms.

6.9 The Periodic Table: Electron Configurations and Valence Electrons

Learning Goals: You should be able to:

1. Describe what we mean by the *s, p, d,* and *f* blocks of elements.
2. Write the electron configuration and valence electron configuration for any element once you know its place in the periodic table.

Electromagnetic Radiation (6.1)

Topic Summaries and Exercises

When you sense the warmth of a fire in a fireplace you are feeling what scientists call **radiation,** or **electromagnetic radiation.** The fire gives off light (visible radiation) and heat (thermal radiation). Both types of electromagnetic radiation exist in the form of **electromagnetic waves.**

- Note in Figure 6.3 in the text the following characteristics of electromagnetic waves: wavelength (λ) and amplitude.
- Electromagnetic waves move through a vacuum at 3×10^8 m/s, the "speed of light."
- Electromagnetic waves differ from one another by their **frequency** (v). This is the number of cycles per second (1 hertz = 1 cycle/s).

Chemists often characterize electromagnetic radiation by its **wavelength,** λ.

- λ measures the distance between two adjacent maxima (peaks) in a periodic wave.
- A key relationship between wavelength, λ, and frequency is:

$$v = \frac{c}{\lambda}$$

where c is the speed of light, 3.00×10^8 m/s.

- Examples of wavelength units are in the following table

Unit	Symbol	Length
Nanometer	nm	1×10^{-9} m
Angstrom	Å	1×10^{-10} m
Picometer	pm	1×10^{-12} m

EXERCISE 1 Characterizing the properties of a wave

Figure 6.1 shows two periodic waves, 1 and 2. (**a**) What do the distances *a* and *b* correspond to? (**b**) Which electromagnetic radiation, 1 or 2, has the greater frequency?

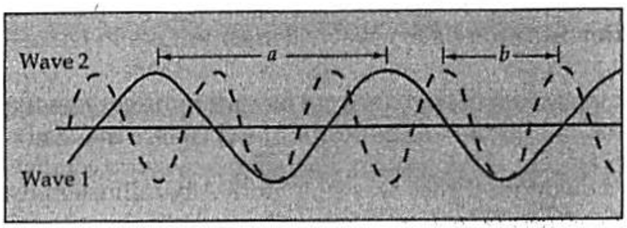


FIGURE 6.1 Two periodic waves

SOLUTION: (a) The distance *a* corresponds to the wavelength for wave 1, and similarly, the distance *b* corresponds to the wavelength for 2. **(b)** Frequency is defined as the number of times per second that maxima of a periodic wave pass a given point. We can see by inspection of Figure 6.1 that the adjacent maxima of wave 2 are closer together than those of wave 1. Because of this closer proximity of adjacent maxima, wave 2 will have more maxima passing a given point in one second than will wave 1. Thus wave 2 has the higher frequency.

EXERCISE 2 Converting units of wavelength

Convert 2×10^{-10} m to nanometer.

SOLUTION: *Analyze*: We are asked to convert 2×10^{-10} m to nanometer.

Plan: We will do the following conversion: *meter* → *nanometers*. The conversion is done with the appropriate numerical value of the prefix *nano-*.

Solve: The unit conversion is made as follows: A nanometer is equal to 1×10^{-9} meter; therefore, the conversion from meter to nanometer is

$$\text{Nanometers} = (2 \times 10^{-10}\,\text{m})\left(\frac{1\,\text{nm}}{1 \times 10^{-9}\,\text{m}}\right) = 2 \times 10^{-1}\,\text{nm}$$

Check: A check of the units show that they cancel to give the desired unit.

EXERCISE 3 Determining wavelength from frequency of electromagnetic radiation

Calculate the wavelength of electromagnetic radiation that has a frequency of 9.22×10^{17}/s.

SOLUTION: *Analyze*: We are asked to calculate a wavelength of electromagnetic energy given its frequency.

Plan: The relationship between frequency and wavelength of electromagnetic radiation is:

$$v = \frac{c}{\lambda}$$

where v has units of s^{-1} (or /s). We want to calculate the wavelength, λ. It is convenient to change the equation so the term λ is on the left-hand side of the equation.

Solve: Rearranging the equation and substituting the known values of c and v gives:

$$\lambda = \frac{c}{v} = \frac{3.00 \times 10^{8}\,\text{m/s}}{9.22 \times 10^{17}/\text{s}} = 3.25 \times 10^{-10}\,\text{m}$$

Quantization of Energy (6.2)

During the nineteenth century, theories of classical physics were unable to explain certain physical phenomena. For example, the color of an extremely hot iron bar was predicted to be blue; yet it is red. Max Planck proposed that light, which had been previously thought to consist of a continuous collection of electromagnetic waves, consisted, instead, of bundles of energy.

- Each bundle of energy has a discrete (individually distinct or separate) value.
- The smallest allowed increment of energy gained or lost is called a **quantum**.
- *Planck's relationship* between the smallest amount of energy gained or lost and the frequency of the associated radiation is

$$E = hv$$

$$h = 6.626 \times 10^{-34}\,\text{J-s (Planck's constant)}$$

Acceptance of Planck's quantum theory was slow until Albert Einstein used it successfully in 1905 to explain the **photoelectric effect**. This effect occurs when a light shines on the surface of a clean metal and electrons are ejected from the surface. Einstein proposed that a beam of light consists of a collection of small particles of energy, called **photons.** A photon interacts with an electron on a metal surface and transfers its energy to the electron. An electron is ejected if it gains sufficient energy from the photon to overcome the forces binding the electron to the surface. Any energy from the photon in excess of the binding energy appears as kinetic energy of the emitted electron.

EXERCISE 4 Determining the energy of a photon I

(a) Does the energy of a photon increase or decrease with increasing wavelength?
(b) In Exercise 1, which electromagnetic wave, 1 or 2, has the greater energy?

SOLUTION: *Analyze*: Question **(a)** asks us to describe how the energy of a photon depends on wavelength. We will have to look up or find a relationship between the two. Question **(b)** asks us to apply the information we learned in question **(a)** to two waves described in Exercise 1.

Plan: **(a)** The relation between energy and wavelength is derived from the following two equations:

$$E = hv \qquad \text{and} \qquad v = \frac{c}{\lambda}$$

Substituting c/λ for v in Planck's equation yields

$$E = hv = h\frac{c}{\lambda}$$

(b) In Exercise 1 we determined which wave had the greater frequency. We can use the relationship above to determine which has the greater energy.

Solve: **(a)** By inspection of the derived relationship we see that the energy of a photon is *inversely* proportional to its wavelength; thus, its energy decreases with increasing wavelength. **(b)** In Exercise 2 we found that wave 2 has the greater frequency; since energy is *directly* proportional to frequency, wave 2 has the greater energy.

EXERCISE 5 Determining the energy of a photon II

Calculate the energy of an X-ray photon with a wavelength of 3.00×10^{-10} m.

SOLUTION: *Analyze*: We are asked to calculate the energy of an X-ray photon with a specified wavelength. The information in the previous Exercise is useful for doing this problem.

Plan: Use the relationship $E = \frac{hc}{\lambda}$ to solve for energy.

Solve: Substitute the given or known values of h, c, and λ into this relationship:

$$E = \frac{hc}{\lambda} = \frac{(6.63 \times 10^{-34}\ \text{J-sec})(3.00 \times 10^{8}\ \text{m/sec})}{3.00 \times 10^{-10}\ \text{m}} = 6.63 \times 10^{-16}\ \text{J}$$

Line Spectra and the Bohr Model (6.3)

In the 1800s, Josef Fräunhofer and others had shown that light emitted from the sun, or from hydrogen atoms heated in a partial vacuum, is not a continuum.

- A continuum is a collection of electromagnetic waves of all possible wavelengths in a given energy region.
- A prism separates light emitted from the sun, or from heated atoms, into discrete and noncontinuous wavelength components.
- A collection of discrete lines or wavelengths on a photographic plate is called a **line spectrum.** See Figure 6.11 in the text.
- Each gaseous monatomic element has its own characteristic line spectra.

In 1914, Niels Bohr proposed a model for the hydrogen atom that enabled scientists to explain hydrogen line spectra. Key ideas in **Bohr's model** of the hydrogen atom are:

- An electron moves in a circular path, called an **orbit**.
- Only orbits restricted to certain energies and radii are permitted.
- Each allowed orbit is assigned an integer n, known as the **principal quantum number**:

$$n = \underbrace{1, 2, 3, 4, \ldots, \infty}_{\text{range of values of } n}$$

- The energy of an orbit in a hydrogen atom is quantized:

$$E_n = -hcR_H\left(\frac{1}{n^2}\right)$$

n = principal quantum number of an orbit

R_H is the Rydberg constant

$-hcR_H = -2.18 \times 10^{-18}\ \text{J}$

- Note that as the principal quantum number increases the energy of an orbit increases (becomes a smaller negative number).
- The orbit of lowest energy, $n = 1$, is called the **ground state**. All others are termed **excited states**.

Emission or absorption of radiant energy from a hydrogen atom occurs when an electron moves from one orbit to another. The energy change is

$$\Delta E = h\nu$$

$$\underbrace{\Delta E}_{\substack{\text{energy difference}\\ \text{between two orbits}}} = -2.18 \times 10^{-18}\ \text{J}\left(\underbrace{\frac{1}{n_f^2}}_{\substack{\text{final orbit of}\\ \text{electron}}} - \underbrace{\frac{1}{n_i^2}}_{\substack{\text{starting (initial)}\\ \text{orbit of electron}}}\right)$$

Emission: electron radius of Bohr orbit decreases: $n_i > n_f$

Absorption: electron radius of Bohr orbit increases: $n_f > n_i$

EXERCISE 6 Identifying the postulates of Bohr's model of a hydrogen atom

What are the postulates of Niels Bohr's model for the hydrogen atom?

SOLUTION: The postulates are: **(a)** An electron moves in a circular path about the nucleus and does not collapse into the nucleus. This movement about the nucleus is characterized by the electron's orbit. **(b)** The energy of an electron can have only certain allowed values; its energy is quantized. **(c)** An electron can move from one circular path to another circular path only when it absorbs or emits a photon corresponding to the exact energy difference between the two quantized energy states.

EXERCISE 7 Meaning of sign of E_n and ionization

What is the meaning of the negative sign in the equation $E_n = -hcR_H(1/n^2)$? What is an excited state? When an electron is ionized, what is its final principal quantum number?

SOLUTION: The negative sign means that energy is required to remove the electron from the nucleus—that is, when the electron and nucleus are combined, energy is released. An alternative explanation is that the electron and nucleus in the atom are more stable than the separated electron and nucleus. Any orbit of the hydrogen atom with $n \geq 2$ is considered a higher energy orbit and is said to be an "excited state." Ionization corresponds to a transition to a final state of $n = \infty$.

EXERCISE 8 Determining energy change when an electron changes its orbit

An electron moves from a $n = 1$ Bohr hydrogen orbit to a $n = 2$ Bohr hydrogen orbit. **(a)** What is the energy associated with this transition? **(b)** Is this electronic transition accompanied by the emission or absorption of energy?

SOLUTION: *Analyze:* In the first part of the question we are asked to determine the energy associated with an electron moving from the $n = 1$ Bohr orbit to the $n = 2$ Bohr orbit. We will need a relationship between energy and the principal quantum number of a Bohr orbit. The second part asks us to determine whether energy is absorbed or emitted when an electron changes from a lower numbered Bohr orbit to a higher one.

Plan: **(a)** The energy change associated with an electron moving from one Bohr orbit to another is given by the relation:

$$\Delta E = (-2.18 \times 10^{-18}\,\text{J})\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right)$$

where n_i is the principal quantum number of the initial orbit and n_f is the principal quantum number of the final orbit. **(b)** The sign of ΔE tells us whether energy is absorbed (positive sign) or emitted (negative sign).

Solve: **(a)** Solve for ΔE by substituting into the previous equation the values $n_i = 1$ (the initial orbit) and $n_f = 2$ (the final orbit):

$$\Delta E = (-2.18 \times 10^{-18}\,\text{J})\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right) = (-2.18 \times 10^{-18}\,\text{J})\left(\frac{1}{2^2} - \frac{1}{1^2}\right)$$

$$= (-2.18 \times 10^{-18}\,\text{J})\left(\frac{1}{4} - 1\right) = (-2.18 \times 10^{-18}\,\text{J})\left(-\frac{3}{4}\right) = 1.63 \times 10^{-18}\,\text{J}$$

(b) The sign of the value for ΔE is positive; thus, the transition requires energy. For this transition to occur, a photon with an energy of at least 1.63×10^{-18} J would have to be absorbed by the electron.

Note: If the energy change had been negative then we know emission occurs. If we are asked to calculate frequency or wavelength from the calculated energy we must always use a positive energy as both frequency and wavelength are positive numbers.

Principles of Modern Quantum Theory (6.4–6.6)

As discussed in the previous topic summary sections, electromagnetic radiation has both wave and particle aspects. In 1924, Louis de Broglie postulated that all material particles have wave properties.

- The result of his conjecture is the **de Broglie relation:**

$$\lambda = \frac{h}{p} = \frac{h}{mv}$$

  where λ is the wavelength of the matter wave, h is Planck's constant, p is the momentum of the moving particle (which equals $m \times v$), m is the mass of the particle, and v is the velocity of the particle.
- Note that for large particles, such as a ball, m is very large; therefore λ is extremely small and not measurable. For small particles, such as an electron, λ is larger and measurable.

Werner Heisenberg proposed that we must view the properties of position and momentum for subatomic particles differently than for macroscopic forms of matter. **Heisenberg's uncertainty principle** says that for a subatomic particle we cannot measure exactly in a given instance both its location in space and its momentum (also velocity). Equation 6.9 in the text is a quantitative statement of this principle.

Application of wave-equation principles, the de Broglie postulate, and the uncertainty principle to the properties of an electron moving about a hydrogen nucleus eventually led to the derivation of mathematical functions (given the symbol ψ) that describe the wave behavior of electrons.

- See Figure 6.16 in the text. Large values of ψ^2 represent a high probability of finding an electron at a particular point in space. This density of dots is referred to as an electron-density distribution.
- An **orbital** is a one-electron wave function and has an associated allowed energy state. An orbital also is represented as an electron-density distribution in space.

Each hydrogen-like orbital is described by three characteristic **quantum numbers**: n, l, and m_l.

- All three are required to completely characterize a hydrogen-like orbital. Their names and their relationships to one another are summarized in Table 6.1 on the next page.
- A collection of orbitals having the same value of n is called an **electron shell.**
- A **subshell** consists of all orbitals with the same n and l values.
- For a hydrogen atom all orbitals with the same n values have the same energy. This is not true for multi-electron atoms.

TABLE 6.1 Relationship among Quantum Numbers n, l, and m_l

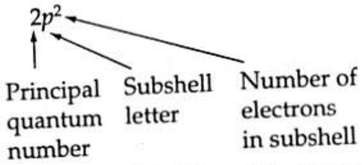
Quantum	Name number	Dependence on other quantum numbers	Range of values
n	Principal	None	$1, 2, 3, \ldots, \infty$
l	Angular momentum	Integral values that depend on the value of n. For each n quantum number, there are n number of l values.	$0, 1, 2, \ldots, n - 1$ for each n value
m_l	Magnetic	Integral values that depend on the value of l. For each l quantum number, there are $2l + 1$ possible m_l values.	$l, l - 1, l - 2, \ldots 0, \ldots, -(l - 1), -l$ for each l value

A shorthand notation is used to describe each hydrogen-like orbital in terms of its n and l quantum numbers.

- In this notation system the principle quantum number, n, appears in front of the l value for the orbital, where the value of l is designated by a letter:

Quantum number l	0	1	2	3
Designation of orbital	s	p	d	f

For example, the shorthand notation for a hydrogen-like orbital with quantum numbers $n = 2$ and $l = 1$ and having two electrons is



Note that the m_l value of a hydrogen-like orbital is not indicated.

- Note that all hydrogen-like orbitals with the same n and l values, such as the $2p$ orbitals, form a subshell.
- Remember that each n level contains n^2 orbitals; each l subshell level contains $2l + 1$ orbitals. For example, the $n = 2$ principle quantum level contains a total of 2^2 or 4 orbitals: One $2s$ orbital and three $2p$ orbitals.

Graphical representations of the $1s$, $2s$, $2p$, and $3d$ orbitals provide you with an understanding of the properties of orbitals.

- Be able to reproduce Figures 6.19, 6.21, 6.22, and 6.23 from your text.
- When looking at plots of electron density such as ψ^2 it is important to remember that we cannot know accurately and simultaneously both the position of an electron and its momentum as stated by Heisenberg's uncertainty principle. Since one cannot measure simultaneously all properties of an electron, we can only talk about its probability of being at a particular point in space. A common type of graphical representation is a 90-percent boundary plot, such as the one shown in Figure 6.2.

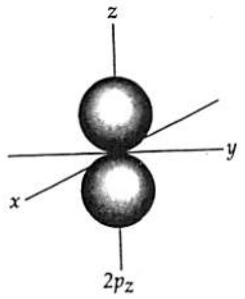


FIGURE 6.2 A 90-percent boundary plot for a $2p_z$ orbital. The spherical surfaces enclose 90 percent of the total ψ^2 for the orbital. Note that there is no electron density in the xy plane. Such a surface is called a *node*.

EXERCISE 9 Identifying the significance of n, l, and m_l

Each wave function (ψ) describing an orbital has three quantum numbers, n, l, and m_l, associated with it. What do these quantum numbers determine in terms of orbital properties?

SOLUTION: The principal quantum number, n, defines the size of an orbital and its energy. As the value of n increases so does the energy of the orbital with respect to the ground-state energy. That is, the energy of an orbital becomes a smaller negative quantity as the value of n increases. The shape of an orbital is related to its angular momentum quantum number, *l*. For a given *l* value for an orbital, the magnetic quantum number, m_l, determines the orientation of the orbital in space with respect to an axis in the presence of a magnetic field.

EXERCISE 10 Determining the number of orbitals with the same *n* value

How many orbitals exist with the same $n = 3$ value but with differing *l* and m_l values?

SOLUTION: *Analyze*: We are asked to determine the total number of orbitals that can have the same *n* value (3) but with differing *l* and m_l values. Thus, we need to know the relationships among *n*, *l*, and m_l.

Plan: Orbitals are described by the quantum numbers *n*, *l*, and m_l: All three are needed to describe a single orbital. Start with $n = 3$ and determine how many *l* states are possible. Then, for each *l* state, determine how many m_l states are possible. Each combination of the three quantum numbers describes an orbital. Finally, sum the number of orbitals that are described to obtain the total.

Solve: For $n = 3$ there exists orbitals with angular momentum quantum numbers $l = 0, 1, \ldots, n - 1$. For each *l* value, there exists orbitals with the following range of m_l values: $l, (l - 1), \ldots 0, \ldots, -(l - 1), -l$. These orbitals are tabulated as follows:

l	m_l	Number of orbitals
0	0	1
1	1, 0, −1	3
2	2, 1, 0, −1, −2	5

The total number of orbitals that can have an $n = 3$ value is therefore 9.

Check: The total number of for a given *n* level equals n^2. 3^2 orbitals $= 9$ orbitals which confirms the analysis in the solution.

EXERCISE 11 Writing shorthand notations for orbitals

Write the shorthand notation (for example, 2*s*) for the subshells described by the following quantum numbers: **(a)** $n = 2, l = 1, m_l = 0$; **(b)** $n = 3, l = 2, m_l = 1$; **(c)** $n = 4, l = 3, m_l = 2$.

SOLUTION: *Analyze*: We are given three sets of *n*, *l*, and m_l values and asked to write the shorthand notation for each subshell.

Plan: A shorthand notation for a subshell has the format nl^x. The value of the principal quantum number, *n*, is listed first. A letter is used in place of the value of *l* (for example, *s* for $l = 0$ and *p* for $l = 1$). The value of m_l is not indicated for a subshell.

Solve: For each case, determine the letter designation for the given *l* value. **(a)** $l = 1$ corresponds to a *p* orbital; thus the subshell designation is 2*p*. **(b)** $l = 2$ corresponds to a *d* orbital; thus the subshell designation is 3*d*. **(c)** $l = 3$ corresponds to a *f* orbital; thus the subshell designation is 4*f*.

EXERCISE 12 Identifying a correct set of quantum numbers for an orbital

Why is it not possible for an orbital to have the quantum numbers $n = 3, l = 2, m_l = 3$?

SOLUTION: *Analyze*: We are asked to explain why a particular set of quantum numbers for a orbital is not possible.

Plan: This type of question requires us to know the relationships among *n*, *l*, and m_l. Start with *n* and determine if the given value of *l* is possible; repeat this for *l* and m_l.

Solve: The possible *l* values for $n = 3$ are $l = 0, 1, 2$; thus the values of $n = 3$ and $l = 2$ are permitted. For a given *l* value, only certain m_l values are allowed: $m_l = \{-l \ldots 0 \ldots +l\}$. The maximum m_l value for $l = 2$ is 2. Thus, $m_l = 3$ is not possible for a orbital with a $l = 2$ value.

Check: The value of m_l can never be larger than the associated value of *l*; this is not the case in the question.

EXERCISE 13 Determining the number of orbitals from a shorthand notation for a subshell

How many orbitals are possible for subshells designated by the following shorthand notations: **(a)** 2*s*; **(b)** 3*d*?

SOLUTION: *Analyze*: We are given shorthand for two subshells and asked to determine how many orbitals are possible for each.

Plan: A shorthand notation for a subshell describes a set of orbitals with the same n and l values. To determine the maximum number of orbitals within that set we must determine the possible m_l values. The number of m_l values for a subshell is $2l + 1$.

Solve: **(a)** The letter *s* corresponds to an orbital with $l = 0$. For $l = 0$, there is only *one* possible orbital with $m_l = 0$. **(b)** The letter *d* corresponds to an orbital with $l = 2$. For $l = 2$, there are *five* orbitals with five different orientations in space determined by the m_l values, which equal 2, 1, 0, −1, and −2.

Check: The total number of orbitals associated with a given nl^x shorthand notation is only determined by the value of l: $2l + 1$. Use of this relationship confirms the number of orbitals determined for each shorthand description of orbitals.

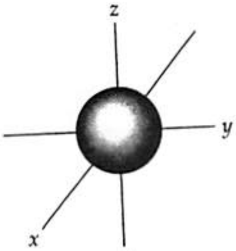
EXERCISE 14 Characterizing a nodal surface

(a) What is a nodal surface? (b) What quantum number determines the number of nodal surfaces?

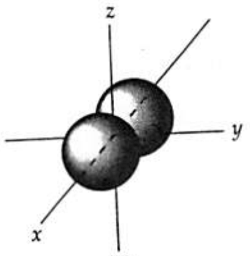
SOLUTION: (a) A nodal surface is a region in space where ψ^2 goes to zero. (b) The number of nodal surfaces, or nodes, depends on the principal quantum number, n. For an orbital with a given n value, there are $n - 1$ nodal surfaces. Thus a $2p$ orbital has one node; a $3d$ orbital has two nodes; and a $1s$ orbital has no nodes. For example, look at Figure 6.2 in the *Student's Guide*. The xy plane is a nodal surface for a $2p_z$ orbital.

EXERCISE 15 Identifying orbitals from their contour representations

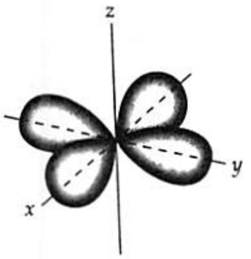
Label the following orbitals as to type:



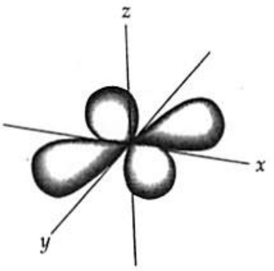
(a)



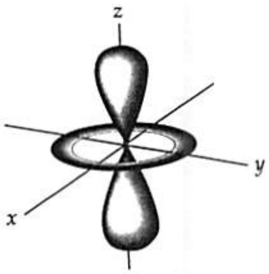
(b)



(c)



(d)



(e)

SOLUTION: *Analyze*: We are given pictures of different orbitals and asked to label them. By label we mean *s* or p_x, for example.

Plan: We either have to find similar orbital pictures in the text or we have memorized the various types and orientations with their labels.

Solve: **(a)** *s*; **(b)** p_x; **(c)** $d_{x^2-y^2}$; **(d)** d_{xy}; **(e)** d_{z^2}.

Check: A review of Figures 6.19, 6.22, and 6.23 in the text confirms the labels.

Energies of Orbitals in Many-Electron Atoms (6.7)

The electronic structure of the hydrogen atom forms the basis of electronic structures for atoms with two or more electrons.

- Orbitals are described like those for hydrogen and have similar shapes.
- We continue to use the orbital designations such as 1*s*, 2*s*, $2p_x$, $2p_y$, and $3d_{xy}$.

In the 1920s, S. A. Goudsmit and B. E. Uhlenbeck proposed that electrons possess intrinsic spin as well as orbital motion.

- Electrons in an atom show two intrinsic spin orientations in the presence of a magnetic field. The two spin orientations are in opposite directions and are quantized.
- Electron spin is quantized. It is defined by the **electron-spin quantum number**, m_s, which has two possible values, $m_s = +\frac{1}{2}$ and $m_s = -\frac{1}{2}$.

There is a limitation placed on quantum numbers for electrons in a many-electron atom. In 1924, Wolfgang Pauli proposed a principle which is known as the **Pauli exclusion principle**.

- According to this principle, *no two electrons in an atom can have the same set of four quantum numbers, n, l, m_l, and m_s.*
- The effect of the Pauli exclusion principle is that an orbital can hold a maximum of two electrons and these electrons must have opposite spins.

The order of energies of orbitals in many-electron atoms is shown in Figure 6.24 in the text. You can use Figure 6.3 by following each arrow from tail to head starting with the 1s orbital and then going to the next arrow in sequence.

- Note that subshells with the same n value (for example, 2s and 2p) do *not* have the same energies as they do in a hydrogen atom.
- Note that orbitals *within* a subshell (for example, $2p_x, 2p_y, 2p_z$) possess the same energy.
- Remember that this ordering of energies is only used to "construct" the electronic structure of atoms.

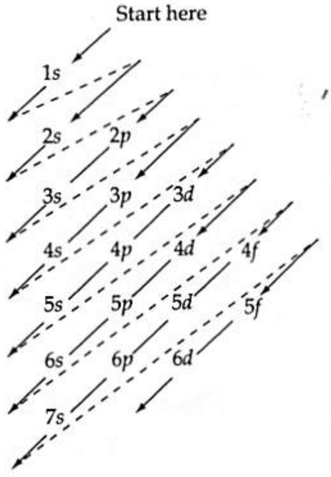


FIGURE 6.3 A technique to aid memorization of the order of orbital energy levels in many-electron atoms. Each horizontal row in the diagram contains all the orbitals for a given n quantum number. The rows follow a sequential listing of n quantum numbers starting from $n = 1$. Parallel arrows are then drawn as shown.

EXERCISE 16 Relating quantum numbers of orbitals to energy

In a many-electron atom, which of the following orbitals, described by their three quantum numbers, have the same energy in the absence of magnetic and electric fields: **(a)** $n = 1, l = 0, m_l = 0$; **(b)** $n = 2, l = 1, m_l = 1$; **(c)** $n = 2, l = 0, m_l = 0$; **(d)** $n = 3, l = 2, m_l = 1$; **(e)** $n = 3, l = 2, m_l = 0$?

SOLUTION: *Analyze*: We are given a set of quantum numbers for five different orbitals and asked to determine which of the five have the same energy when electric or magnetic fields are not present.

Plan: Orbitals with the same n and l values have the same energy in many-electron atoms when electric or magnetic fields are not present.

Solve: Orbitals **(d)** and **(e)**, which are 3*d* orbitals, are the only orbitals with the same n and l values and thus the same energy.

Comment: Note that the question is restrictive in that magnetic and electric fields are not present. If they are present, then the rule that orbitals with the same n and l values have the same energy does not necessarily apply.

EXERCISE 17 Determining the maximum number of electrons occupying a subshell

What is the maximum number of electrons that can occupy each of the following subshells: 1*s*, 2*p*, 3*d*, and 4*f*?

SOLUTION: *Analyze*: We are given four different subshells indicated by their shorthand notations and asked to determine what is the maximum number of electrons that can occupy each.

Plan: A subshell consists of one or more orbitals; thus we first must determine the number of orbitals within each subshell. The number of orbitals within a subshell equals the total number of m_l values for the l value of the subshell, $2l + 1$. The second step is to recognize that a single orbital can hold a maximum of two electrons of opposite spin. Thus, the total number of electrons possible within a subshell is $2(2l + 1)$.

Solve: The maximum number of electrons that can occupy the orbitals in each subshell is below:

Subshell	l value	m_l values	Number of m_l values	Maximum number of electrons in subshell
1*s*	0	0	1	$2 \times 1 = 2$
2*p*	1	$1, 0, -1$	3	$2 \times 3 = 6$
3*d*	2	$2, 1, 0, -1, -2$	5	$2 \times 5 = 10$
4*f*	3	$3, 2, 1, 0, -1, -2, -3$	7	$2 \times 7 = 14$

Check: A review of the maximum number of electrons for each subshell in the problem shows that the answer fits the general formula: maximum number of electrons in a subshell $= 2(2l + 1)$.

Electronic Structure of Many-Electron Atoms (6.8)

At this point, we need to review the rules governing the possible values of the four quantum numbers for electrons:

1. The principal quantum number, n, has integral values of $1, 2, 3, \ldots$
2. The integral values of the l quantum number for a given n are $l = 0, 1, 2, 3, \ldots, n - 1$.
3. The integral values of the m_l quantum number for a given l are $m_l = l, l - 1, l - 2, \ldots, 0, \ldots, -l + 1, -l$.
4. The values of the m_s quantum number for a given m_l are $m_s = +\frac{1}{2}, -\frac{1}{2}$.

With this information we can now specify the four quantum numbers for all electrons in an atom. This description of how electrons are arranged in orbitals is called an **electron configuration**. The procedure for writing the electron configuration of an element is summarized as follows:

1. Electrons occupy orbitals in order of increasing energy. All orbitals that are equal in energy (degenerate) are filled with electrons first, before the next level of energy begins to fill up. (A few exceptions exist, for example, Cr and Cu.)
2. An orbital can have a maximum of two electrons with *opposite* spins. This is an application of the Pauli exclusion principle.
3. When several orbitals of equal energy exist (for example, $2p_x, 2p_y, 2p_z$), the filling of orbitals with electrons follows **Hund's rule**: *Electrons enter orbitals of equal energy singly and with the same spins, until all orbitals have one electron each*. Then electrons with opposite spins pair with the electrons in the half-filled set of equal-energy orbitals.

Three methods are used to describe the electron configuration for an element:

1. Assignment of the four quantum numbers (n, l, m_l, m_s) for each electron;
2. A shorthand using the orbital system notation described in Section 6.5 of the text with the number of electrons indicated (as a superscript) for each set of degenerate orbitals;
3. An orbital diagram (see Exercise 19).

These methods are described in Exercises 18 through 20.

EXERCISE 18 Describing electron configurations for atoms I

Using the four quantum numbers, describe the electron configuration for each of the following atoms: **(a)** B; **(b)** N. Assume that an electron first enters an orbital in the spin state described by the spin quantum number $m_s = +\frac{1}{2}$.

SOLUTION: *Analyze*: We are asked to write the complete electron configurations for B and N using the four quantum numbers.

Plan: First determine the number of electrons possessed by each atom. Then using the *Aufbau Principle, Hund's rule,* and *Pauli's Principle* construct a table that shows the four quantum numbers for each electron.

Solve: (**a**) Boron has five electrons. The first two electrons enter the lowest energy orbital, 1*s*, with opposite spins as per the Pauli exclusion principle. The next energy level, 2*s*, also has two electrons with opposite spins. The last electron enters the 2*p* energy level. The electron configuration for boron is as follows:

n	l	m_l	m_s	**Orbital notation**
1	0	0	$+\frac{1}{2}$	1*s*
1	0	0	$-\frac{1}{2}$	1*s*
2	0	0	$+\frac{1}{2}$	2*s*
2	0	0	$-\frac{1}{2}$	2*s*
2	1	1	$+\frac{1}{2}$	2*p*

Note: The value of m_l for the last orbital could also be written as 0 or −1.

(**b**) Nitrogen has seven electrons. The first five electrons enter orbitals as for boron. The last two electrons enter 2*p* orbitals because they have a maximum occupancy of six electrons. According to Hund's rule, they enter singly until all degenerate orbitals are half filled:

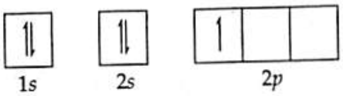
n	l	m_l	m_s	**Orbital notation**
1	0	0	$+\frac{1}{2}$	1*s*
1	0	0	$-\frac{1}{2}$	1*s*
2	0	0	$+\frac{1}{2}$	2*s*
2	0	0	$-\frac{1}{2}$	2*s*
2[a]	1	1	$+\frac{1}{2}$	2*p*
2[a]	1	0	$+\frac{1}{2}$	2*p*
2[a]	1	−1	$+\frac{1}{2}$	2*p*

[a] These three 2*p* orbitals are referred to as a half-filled set of orbitals.

Check: A check of the total number of electrons in the summary of electron configurations for B and N confirms that five and seven electrons are identified. No two electrons have the same set of quantum numbers. If two electrons had the same set of quantum numbers this would tell us that we made an incorrect assignment.

EXERCISE 19 Describing electron configurations for atoms II

Another approach to describing the electron configuration of an atom is an orbital diagram, in which a box represents an orbital. Electrons are represented by arrows, with an arrow that points up (↑) corresponding to the spin state $m_s = +\frac{1}{2}$ and an arrow that points down (↓) corresponding to the spin state $m_s = -\frac{1}{2}$. For example, the electron configuration for boron can be represented as

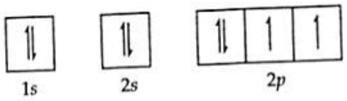


What are the orbital diagrams for: (**a**) O; (**b**) Sr; (**c**) V?

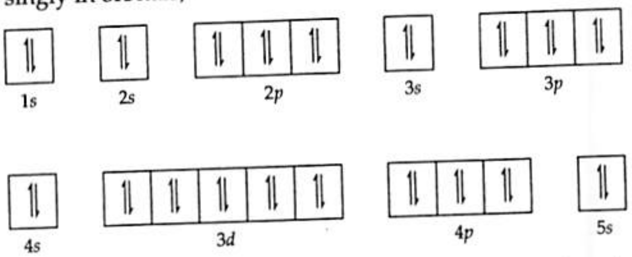
SOLUTION: *Analyze:* We are given information about how to construct electron box diagrams and are asked to construct electron box diagrams for O, Sr, and V.

Plan: First determine the number of electrons possessed by each atom. Then, using a box for each orbital within a subshell, construct the appropriate number of boxes, grouped as electron subshells, that permits a complete orbital description for the electrons in the atom.

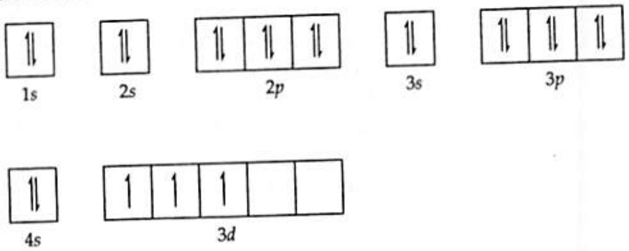
Solve: (**a**) Oxygen has eight electrons:



The electron configuration shown for the degenerate set of $2p$ orbitals conforms to Hund's rule, which requires that electrons remain unpaired until all degenerate (equal energy) orbitals are filled with one electron. Thus, the first three $2p$ electrons are placed singly in orbitals, and the fourth one is paired. (**b**) Strontium has 38 electrons:



When *forming* an electron configuration, fill the 4s orbitals before the 3d orbitals. As a check, you should count the number of electrons. (**c**) Vanadium has 23 electrons:



Check: Each box has no more than two electrons and if two electrons are in the same box they have opposite spins. The order of orbitals is reconfirmed by use of Figure 6.3 in the *Student's Guide*.

EXERCISE 20 Describing electron configurations for atoms III

An orbital diagram is useful and informative, but cumbersome. More commonly, a shorthand notation system is also used to describe electron configurations. This notation uses the *nl* symbol for each subshell (such as 1*s*, 2*s*, 2*p*, 3*s*, 3*d*), with the number of electrons occupying an orbital set indicated with a superscript. For example, boron has five electrons; its electron configuration is $1s^22s^22p^1$. Write the complete shorthand configuration for the atoms in Exercise 19. Also write each using a shorthand for the core electron configuration. For example, [He] represents the $1s^2$ electrons possessed by helium.

SOLUTION: *Analyze*: We are given information on how to write the complete electron configuration for elements and asked to write the complete shorthand notation for all electrons possessed by O, Sr, and V.

Plan: First determine the number of electrons possessed by each atom. Then, use the principles of building electron configurations of atoms to place electrons in appropriate subshells. You have already done this in the previous exercise.

Solve: (**a**) O: $1s^22s^22p^4$; or [He]$2s^22p^4$. (**b**) Sr: $1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^2$; or [Kr]$5s^2$, where [Kr] is used to describe the Kr core electron configuration $1s^22s^22p^63s^23p^64s^23d^{10}4p^6$. (**c**) V: $1s^22s^22p^63s^23p^64s^23d^3$ or [Ar]$4s^23d^3$, where [Ar] is used to describe the Ar core electron configuration $1s^22s^22p^63s^23p^6$.

Check: The number and types of electrons agree with the assignments in Exercise 19. Always count the total number of electrons in the shorthand notation to confirm that it agrees with the atomic number of the element.

The Periodic Table: Electron Configurations and Valence Electrons (6.9)

The electronic configuration of an element can also be constructed from its location in the periodic table.

- Elements in **families** (*vertical columns*) of representative elements possess the same number and type of outer electrons beyond the core electron configuration. These outer-shell electrons are referred to as **valence electrons**. The distribution of valence electrons in transition elements can vary slightly within a family.
- Valence electrons are involved in ion formation and chemical reactions to form new chemical bonds.

- If you know the valence electron configuration for the first member of a representative family of elements you then know the valence electron configuration for the other members. For example, the electron configuration for oxygen is

$$\text{O: }\underset{\substack{\nearrow\\ \text{core}\\ \text{noble-gas}\\ \text{electron}\\ \text{configuration: [He]}}}{1s^2}\quad \underset{\substack{\nwarrow\\ \text{valence electrons}}}{2s^22p^4}$$

Any other member of the oxygen family will have a similar valence electron configuration:

$$\underset{\substack{\nearrow\\ n\text{ represents the period}\\ \text{the element resides in}}}{ns^2}\underset{\substack{\nwarrow\\ \text{note the }s^2p^4\text{ arrangement}\\ \text{of valence electrons, the}\\ \text{same as in an oxygen atom}}}{np^4}$$

A knowledge of the period in which an element resides will help you write complete electron configurations. Note the following key points shown in Figures 6.30 and 6.31 in the text:

- The first three periods consist only of **representative elements**: Only *s* and *p* orbitals are involved in their valence electron configurations. The period number corresponds to the principal quantum number of the valence electrons.
- Periods 4–7 contain both representative elements and **transition elements**. The ns and $(n-1)d$ outer orbitals are occupied in transition elements, where n is the period number. These form the valence orbitals for transition elements. The valence orbitals of transition elements are the ns $(n-1)d$ orbitals. Note that beyond the transition elements in a period, the *d* electrons form part of the inner core of electrons and are no longer valence electrons.
- The **inner-transition elements** have *f* orbitals occupied.

EXERCISE 21 Writing valence-electron configurations for elements

(**a**) Write the valence electron configuration for the elements of the second period. (**b**) How does the group number of a representative family relate to its valence-electron configuration?

SOLUTION: *Analyze*: We are asked to write the valence-electron configurations for the elements of the second row: Li, Be, B, C, N, O, F, and Ne. We are also asked to explain how the group number of the family helps us to construct valence-electron configurations.

Plan: The period number identifies the principal quantum number for the valence electrons in that period. (Note: This is only for the *s* and *p* electrons; if an atom possesses *d* valence electrons then for the *d* valence electrons the principal quantum number is $n-1$.) Starting from the left of the period, and as we move from element to element, we fill the *s* orbital first with electrons and then the *p* orbitals. This period does not contain transition metals and we do not have to use *d* orbitals. For the other two questions we need to look at the valence-electron configuration of each element and compare it to the other elements in the same family.

Solve: (**a**) The second period of elements begins after $H(1s^1)$ and $He(1s^2)$. The next quantum level to be filled is $n = 2$. Thus the valence shell of each element contains electrons in the energy level $n = 2$. Using Hund's rule and the Pauli exclusion principle, we obtain the following valence-electron configurations:

Family:	1A	2A	3A	4A
Element:	Li	Be	B	C
Electron configuration:	$2s^1$	$2s^2$	$2s^22p^1$	$2s^22p^2$
Family:	5A	6A	7A	8A
Element:	N	O	F	Ne
Electron configuration:	$2s^22p^3$	$2s^22p^4$	$2s^22p^5$	$2s^22p^6$

(b) The group number of a representative family equals the number of valence electrons for the family.

Check: The number of electrons in each valence-electron configuration agrees with the group number of the family in which the element resides. The valence electrons are assigned to the highest energy subshell. The period number corresponds to the *n* value in the shorthand notation, two.

EXERCISE 22 Writing complete electron configurations of elements using the periodic table

Use the periodic table to write the complete electron configurations of S and Ca.

SOLUTION: *Analyze*: We are to use the structure of the periodic table to help us write the complete electron configurations of S, a member of the oxygen family, and Ca, a member of the alkaline-earth family.

Plan: First determine the number of electrons possessed by each element from its atomic number. Each period (row) has associated with it a set of valence-shell orbitals used by the atoms within the period. We encountered this situation in the previous exercise. Fill the orbitals with electrons, period-by-period in a sequential manner, until the number of electrons possessed by the element is reached.

Solve: Sulfur is in the 6A [16] family and in the third period; it has 16 electrons. We can complete the electron configuration by writing the valence orbitals used by atoms within each period and adding the electrons, period by period, to fill the orbitals. Adding 16 electrons to valence orbitals requires using the first three periods.

Period 1: $1s^2$

Period 2: $2s^22p^6$

Period 3: $3s^23p^4$ (the 3*p* orbital has only four electrons because the total of 16 electrons has been reached)

Complete electron configuration: $1s^22s^22p^63s^23p^4$

Calcium is in the 2A [2] family and is in the fourth period; it has 20 electrons.

Period 1: $1s^2$

Period 2: $2s^22p^6$

Period 3: $3s^23p^6$

Period 4: $4s^2$ (the total number of electrons is now 20)

Complete electron configuration: $1s^22s^22p^63s^23p^64s^2$.

Check: We can confirm the answers by using procedures in Exercise 20.

EXERCISE 23 Writing valence-electron configurations for transition elements

Write the valence-electron configurations for the following transition elements: Ti, Cr, Cd, and Cu.

SOLUTION: *Analyze*: We are asked to write the valence-electron configurations of four transition elements. The periodic table can guide us in doing the problem.

Plan: We can answer this question by determining the electron configuration of each element and identifying the valence electrons for transition elements by recognizing that they are in the outer $(n - 1)$ *d* and *ns* orbitals.

Solve: Ti: $3d^24s^2$. Ti is in the fourth period and therefore *n* is four.
Cr: $3d^54s^1$. Cr is an exception discussed in the text and the resulting valence-electron configuration has two half-filled subshells.
Cd: $4d^{10}5s^2$. Cd is the last transition element in the fifth period; thus its 4*d* orbitals are completely filled.
Cu: $3d^{10}4s^1$. This situation is similar to that discussed in the answer for Cr. A filled *d* orbital and a half-filled *s* orbital are apparently more stable than $3d^94s^2$.

Self-Test Questions

Having reviewed key terms in Chapter 6, match key terms with phrases and identify statements as true or false. If a statement is false, indicate why it is incorrect.

Match each phrase with the best term:

6.1 In the expression $\Delta E = h\nu$ the name given to the term $h\nu$.
6.2 The name given to the pattern of energy emitted when a gas in placed under reduced pressure and high voltage is passed through it.
6.3 Radiation which is characterized by a wave with amplitude and frequency.
6.4 A spectrum characterized by a non-continuous pattern of energies.
6.5 The statement that tells us that if we know accurately the speed of an electron we cannot measure accurately its position.
6.6 The term given to the collection of 2*s* and 2*p* set of orbitals.
6.7 A term used to describe orbitals that have the same energy.
6.8 Mo is in this cluster of elements.
6.9 Eu is in this cluster of elements.
6.10 Mg and P belong to elements which are given this name.
6.11 It has two possible values, +1/2 or −1/2.
6.12 Particles with energy but no mass.
6.13 A mathematical equation that describes the periodic properties of an electron moving about a hydrogen nucleus.
6.14 A term used to describe the probability of finding an electron in space.
6.15 The allowed wave function for an electron.
6.16 The name given to the condition that a 2*p* hydrogen orbital possesses no electron density at the coordinates $x = y = z = 0$.

Terms:

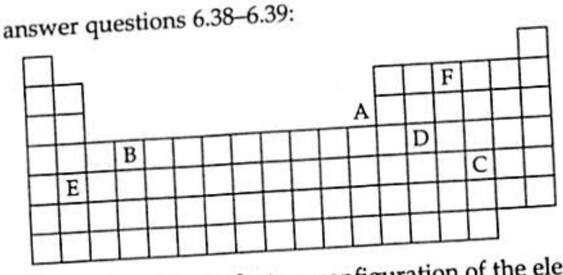
(a) degenerate
(b) electromagnetic
(c) electron density
(d) electron-spin quantum number
(e) Heisenberg's uncertainty principle
(f) lanthanide
(g) line spectrum
(h) node
(i) orbital
(j) photons
(k) representative
(l) spectrum
(m) subshells
(n) transition
(o) quantum of energy
(p) wave function

True-False Statements:

6.17 According to Louis de Broglie's postulate, the wavelength of a *matter wave* associated with an electron increases with the velocity of an electron as it moves about a nucleus.
6.18 The 3*s* and 3*p* orbitals comprise the third *electron shell*.
6.19 Uranium is an *f-block element*.
6.20 The *electron configuration* for carbon is $1s^2 2s^2 2p^3$.
6.21 The *valence* electrons for fluorine are $2p^3$.
6.22 According to the *Pauli exclusion principle*, two electrons in an atom can have the following set of quantum numbers: $n = 2, l = 1, \mathrm{m}_l = 0, m_s = +\frac{1}{2}$ and $n = 2, l = 1, m_l = 0, m_s = +\frac{1}{2}$.
6.23 In titanium, the two 3*d* electrons are predicted to be unpaired by application of *Hund's rule*.
6.24 The *electronic structure* of an atom refers to the number of electrons possessed by it.
6.25 Electromagnetic radiation moves at the speed of light in a vacuum.
6.26 The longer the *wavelength* of radiant energy, the higher the energy.
6.27 Frequency of a wave is directly proportional to wavelength.
6.28 When sunlight is passed through a prism, a *continuous spectrum* of all wavelengths is produced.
6.29 The *ground state* for a helium atom is $1s^1 2s^1$.
6.30 When a 2*s* electron of a carbon atom gains energy and occupies a 3*s* orbital, an *excited state* of carbon is produced.
6.31 According to de Broglie, the greater the *momentum* of a particle, the longer the wavelength of its matter wave.
6.32 All materials possess a *matter wave*.
6.33 The *probability density* for an electron in an atom is related directly to its wave function.
6.34 There are four orientations for *electron spin*.
6.35 *Core electrons* are those electrons in the shell closest to the nucleus.
6.36 The fourth row of the periodic table contains *transition metals*.
6.37 Calcium is an *active metal*.

Problems and Short-Answer Questions

Use the following figure to answer questions 6.38–6.39:



6.38 What is the characteristic ground-state valence-electron configuration of the elements in the family labeled by A?

6.39 Which of the elements identified by the letters B, C, D, E, and F have only four valence electrons? Explain your answer.

6.40 Calculate the energy of a photon with a wavelength of 420 nm.

6.41 Using the Bohr model, calculate the wavelength of radiant energy associated with the transition of an electron in the hydrogen atom from the $n = 4$ orbit to the $n = 2$ orbit. Is energy emitted or absorbed? Indicate the type of radiation emitted.

6.42 One of Bohr's postulates is that an electron moves in a well-defined path about the nucleus. What principle later negated this postulate? Why?

6.43 Differentiate between the functions ψ and ψ^2.

6.44 For each of the following pairs of orbitals of a hydrogen atom, which orbital has the higher energy?

(a) 2*s* or 3*s* **(b)** 3*p* or 4*d* **(c)** 2*s* or 2*p*

6.45 A hydrogen orbital is described by the quantum numbers $n = 4, l = 2$, and $m_l = 0$.

(a) Write the shorthand notation for this orbital.

(b) Are there any other hydrogen orbitals that have the same shorthand notation and if so, which ones?

6.46 Write the shorthand notation for all orbitals of a hydrogen atom that belong to the $n = 4$ principal quantum number energy level. What property do they all have in common?

6.47 Give the values of n, l, and m_l for each orbital in the $n = 2$ principal quantum number energy level.

6.48 An instructor shows you a three-dimensional picture of the surface of an orbital with a spherical shape. Is there sufficient information for you to conclude that the orbital is a 1*s* orbital? If there is insufficient information, what additional information do you need to determine the type of orbital?

6.49 An atom of an element has a valence-electron configuration that includes three 4*p* electrons. Is there sufficient information for you to determine which element it is? If there is insufficient information, what additional information do you need to determine the type of orbital?

6.50 Which orbital notation violates the Pauli Principle and why? **(a)** $2s^2$; **(b)** $3p_x^0$; **(c)** $4d_{xy}^2$; **(d)** $5d_{xz}^3$

6.51 What is a node? Is there a general relationship between the number of nodes possessed by a wave function and its energy? What differences in orbital size and number of nodes do you expect between a 3*s* and a 4*s* hydrogen orbital?

6.52 Draw the contour representations for the $3d_{z^2}$, $3d_{x^2-y^2}$, and $3d_{xy}$ orbitals. In your sketches, not all of the orbitals will look identical. Does this indicate that these 3*d* orbitals are all different in energy?

6.53 Louis de Broglie showed that an electron's wavelength and momentum are related by the equation $\lambda = h/mv$. Why is the relationship $E = hc/\lambda$ not valid for matter typically used by humans.

6.54 Microwave ovens are now commonly used to heat foods. Most microwave ovens have warnings posted to the effect that microwave radiation is harmful to humans. A typical microwave frequency is 20,000 megacycles per second. Calculate the wavelength and energy of this radiation. Show that this energy is significantly less in magnitude than that of visible radiation with a frequency of 100,000,000 megacycles per second and, thus, the danger of microwave radiation cannot be attributed to its energy content.

6.55 **(a)** In what orbitals and in what order are electrons entering the atoms of the second and fifth periods?

(b) The valence orbitals of the elements of the fourth period are the 4*s*, 3*d*, and 4*p* orbitals. Why isn't the fourth period composed of elements with only outer 4*s* and 4*p* orbitals?

6.56 Write the shorthand notation and the box-diagram electron-configuration representation for each of the following elements and indicate the outer orbital(s):

(a) N **(b)** K **(c)** Cu **(d)** Al

6.57 Each of the following electron configurations represents the valence-electron configuration for an atom of an element in the periodic table. For each, what is the element, and in what period does it belong?

(a) $4s^24p^2$ **(c)** $6s^1$ **(e)** $4s^23d^{10}$

(b) $4s^23d^6$ **(d)** $2s^22p^5$

6.58 An atom contains layers or shells of electrons. A completely filled layer or shell of electrons in an atom is a very stable electron configuration.

- **(a)** Both Ne and Xe contain completely filled shells of electrons. What electrons comprise the shell that gets filled between Ne and Ar?
- **(b)** When a representative element forms an ion, it usually forms one whose electron configuration corresponds to the electron configuration for the nearest inert gas. What electrons must be added to or removed from each of the following atoms to form an ion with the nearest inert gas electron configuration: Li; Sr; O?

6.59 Which of the following electron configurations for neutral atoms correspond to ground states, and which correspond to excited states?

(a) $1s^1 2s^1$
(b) $[Ar]4s^1$
(c) $[Ar]3d^6$
(d) $[Xe]6s^2 5d^{10} 6p^3$

6.60 Identify the groups of elements having outer electron configurations of **(a)** ns^2np^4; **(b)** $ns^2(n-1)d^1$; **(c)** ns^2; **(d)** ns^2np^1.

Integrative Questions

6.61 Student (A) is asked to assign a set of n and l values to one of the valence electrons in a certain atom. The student reports that one of the valence electrons has the values of $n = 4$ and $l = 1$. Another student (B) looks at this information and concludes that the element must be arsenic.

- **(a)** Is student (B)'s conclusion reasonable? Explain.
- **(b)** If student (B)'s conclusion is not reasonable what additional information in terms of the electrons would the student need to correctly identify the element?
- **(c)** Student (A) reports that the element is very stable and very unreactive. Does this information help student (B) identify the element? Explain.
- **(d)** Student (A) reports the element is metallic. Does this help student (B) identify the element? Explain.

6.62 Sodium arc lamps emit an intense yellow color resulting from the emission spectrum of sodium atoms. There are two closely-spaced emission lines with a center at 589.3 nm that are primarily responsible for the yellow color.

- **(a)** What is the energy of a photon with that wavelength?
- **(b)** What is the energy emitted by a mole of sodium atoms for that wavelength?
- **(c)** What is the electron configuration of sodium? What are its valence electrons?
- **(d)** What is the electron configuration of Na^+?

6.63 Richard Feynman, a notable physicist, wrote in 1964 in *The Character of Physical Law*, The Messenger Lectures, 1964, MIT Press, pp. 127–128, that "Electrons behave ... in exactly the same ways as photons; they are both screwy, in exactly the same way."

- **(a)** Based on your reading of Chapter 6 why do you think Feynman characterized both electrons and photons as "screwy"?
- **(b)** Feynman also said that our desire for familiarity will lead to perpetual torment as we learn about quantum mechanics. Does this make psychological sense? Thus, what must you do as you study the principles of quantum mechanics to avoid this torment?

Multiple-Choice Questions

6.64 A quantum of electromagnetic radiation has a wavelength equal to 7.52×10^6 Å. What is the frequency of this radiation in cycles/sec?

(a) 1.13×10^{-12}
(b) 8.80×10^{-26}
(c) 3.99×10^{11}
(d) 9.45×10^{13}
(e) 2.69×10^{-15}

6.65 What is the energy of radiation that has a frequency of 9.00×10^{11} cycles/sec? (Remember that Planck's constant, h, has a value of 6.63×10^{-34} J-sec.)

(a) 1.66×10^{-45} J
(b) 5.97×10^{-22} J
(c) 4.99×10^{-27} J
(d) 5.00×10^{-22} J
(e) 3.32×10^{-45} J

6.66 Calculate the wavelength of an electron traveling with a velocity of 4.0×10^9 cm/sec in an electron microscope. The mass of an electron is 9.1×10^{-28} g; $h = 6.63 \times 10^{-34}$ J-sec; and $1\ J = 1\ kg\text{-}m^2/sec^2$.

(a) 0.18 Å
(b) 0.67×10^{-8} cm
(c) 1.5 Å
(d) 1.5×10^8 cm
(e) 1.1×10^{-38} Å

6.67 In the photoelectric effect, in order for an electron to be released from the surface of a clean metal, which one of the following conditions must exist?

- **(a)** The metal must have a low temperature;
- **(b)** the metal must have a high temperature;
- **(c)** the kinetic energy of photons striking the metal's surface must equal that of the emitted electron;
- **(d)** the kinetic energy of photons striking the metal's surface must be less than that of the emitted electrons;
- **(e)** the kinetic energy of photons striking the metal's surface must be greater than or equal to that of the emitted electrons plus the binding energy holding the electron in the metal.

6.68 Passing an electrical charge through argon gas contained in a partially evacuated vessel yields which of the following?

(a) a continuous spectrum; **(b)** a line spectrum; **(c)** white light; **(d)** no visible change; **(e)** (a) and (c).

6.69 According to Bohr's model of the atom, which of the following characteristics of metallic elements explains the fact that many of these elements can easily form positively charged ions?

(a) few electrons;
(b) many electrons;
(c) electrons in lowest-energy n level having high energies;
(d) electrons having momentum;
(e) relatively small ionization energies in their ground states.

6.70 According to the Bohr model of the atom, emission of electromagnetic radiation by heated atoms in a vacuum is directly due to which of the following?

(a) photons absorbed by atoms;
(b) particle emission from the nucleus;
(c) momentum possessed by electrons;
(d) electrons being excited from an inner to outer orbit;
(e) electrons falling from an outer to inner orbit.

6.71 Which of the following is *not* true about the Bohr model of the hydrogen atom?

(a) Electrons decay into the nucleus;
(b) the model cannot account for the ionization of an electron;
(c) an electron in a stable Bohr orbit does not emit radiation continuously;
(d) an electron may remain in an orbit indefinitely;
(e) the hydrogen atom absorbs radiant energy in multiples of $h\nu$.

6.72 Which of the following statements is *not* true about the principal quantum number, n?

(a) It is related to the spin of an electron.
(b) It is related to the energy of an electron.
(c) The larger the n value of an electron, the higher its energy.
(d) The larger the n value of an electron, the larger the value of its Bohr radius.
(e) The lowest energy state for an electron in a Bohr atom corresponds to $n = 1$.

6.73 For an orbital with $l = 4$, what are the possible numerical values of m_l?

(a) $1, 0, -1$
(b) $3, 2, 0, -1, -2, -3$
(c) $3, 2, 1, 0$
(d) $4, 3, 2, 1, 0, -1, -2, -3, -4$
(e) $4, 3, 2, 1, 0$

6.74 Which of the following combinations of quantum numbers for an electron is *not* permissible?

(a) $n = 5, l = 2, m_l = 0$
(b) $n = 3, l = 2, m_l = 3$
(c) $n = 4, l = 3, m_l = -2$
(d) $n = 1, l = 0, m_l = 0$
(e) $n = 2, l = 1, m_l = -1$

6.75 Which of the following statements is the most correct and complete with reference to an electron transition from a Bohr $n = 2$ to $n = 4$ hydrogen orbit?

(a) Energy is released during the electronic transition.
(b) Energy is absorbed during the electronic transition.
(c) The energy of the transition is proportional to $\frac{3}{16}$.
(d) Both (b) and (c).

6.76 Which set of hydrogen-like orbitals can contain a maximum of ten electrons?

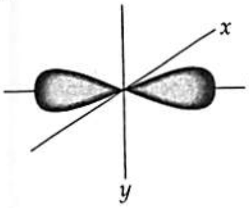
(a) $3s$
(b) $3p$
(c) $4p$
(d) $4d$
(e) $4f$

6.77 For the principal quantum number 3 what are the allowed values of the l quantum number?

(a) $1, 2$
(b) $0, 1, 2$
(c) $1, 2, 3$
(d) $1, 2, 3, 4$
(e) $-2, -1, 0, 1, 2$

6.78 What designation is used to describe the following orbital?

(a) s
(b) p_x
(c) p_z
(d) d_{z^2}
(e) d_{xz}



6.79 In iron, which of the following electrons, characterized by the four quantum numbers, has the lowest energy?

(a) $n = 4, l = 0, m_l = 0, m_s = +\frac{1}{2}$
(b) $n = 3, l = 2, m_l = 1, m_s = -\frac{1}{2}$
(c) $n = 3, l = 2, m_l = 0, m_s = +\frac{1}{2}$
(d) $n = 3, l = 1, m_l = 0, m_s = +\frac{1}{2}$
(e) Electrons (b), (c), and (d) are all degenerate and lowest in energy.

6.80 What is the electron configuration of the element Xe?

(a) $1s^22s^22p^63s^23p^6$
(b) $1s^22s^22p^63s^23p^64s^13d^{10}$
(c) $1s^22s^22p^63s^23p^64s^23d^{10}4p^6$
(d) $1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^2$
(e) none of the above

6.81 How many electrons are there in the outermost shell of arsenic?
(a) 3
(b) 4
(c) 5
(d) 6
(e) 7

6.82 Which element possesses the electron configuration $1s^22s^22p^63s^23p^64s^23d^6$?
(a) V
(b) Cr
(c) Mn
(d) Ni
(e) Fe

6.83 How many electrons in an atom can possess a $n = 3$ principal quantum number?
(a) 3
(b) 6
(c) 8
(d) 10
(e) 18

6.84 Which of the following is the correct order of increasing energy of atomic orbitals when forming a many-electron atom?
(a) $1s2s2p3s3d$
(b) $1s2s3s3p4s$
(c) $1s2s2p3s3p$
(d) $1s2s2p3d4s$
(e) none of the above

Self-Test Solutions

6.1 (o). **6.2** (l). **6.3** (b). **6.4** (g). **6.5** (e). **6.6** (m). **6.7** (a). **6.8** (n). **6.9** (f).
6.10 (k). **6.11** (d). **6.12** (j). **6.13** (p). **6.14** (c). **6.15** (i). **6.16** (h).
6.17 False. $\lambda = h/mv$. The value of λ decreases with increasing velocity.
6.18 False. An electron shell consists of a complete collection of orbitals that have the same value of n. The $3s$, $3p$, and $3d$ comprise the third electron shell.
6.19 True. **6.20** False. The electron configuration for carbon is $1s^22s^22p^2$.
6.21 False. The valence-shell electrons for fluorine are $2s^22p^5$.
6.22 False. They must have different values for m_s because their values for n, l, and m_l are the same.
6.23 True. **6.24** False. Not only does it include the number of electrons, but also their energies.
6.25 True.
6.26 False. Energy and wavelength are inversely related. The longer the wavelength, the smaller the energy.
6.27 False. Frequency and wavelength are inversely related.
6.28 True. **6.29** False. It is $1s^2$. **6.30** True.
6.31 False. The wavelength associated with matter is inversely proportional to momentum. **6.32** True.
6.33 False. Probability density is proportional to the square of the wave function.
6.34 False. There are two spin orientations.
6.35 False. Core electrons are those that are not valence electrons.
6.36 True. **6.37** True. Group IA and IIA elements are active metals.
6.38 The column of elements belongs to the last group of the transition elements. Thus, the valence-shell orbitals are completely filled: $(n-1)d^{10}ns^2$.
6.39 Elements with the ground-state valence electron configuration ns^2np^2 or $(n-1)d^2ns^2$ have a total of four valence electrons. Only elements B and D meet one of these two requirements.
6.40 The relation between the energy of a photon and its wavelength is $E = hc/\lambda$. Substituting the appropriate values for h, c, and the wavelength (converted into the unit meter) yields

$$E = \frac{(6.63 \times 10^{-34}\,\text{J-sec})(3.00 \times 10^{8}\,\text{m/sec})}{(4.20 \times 10^{2}\,\text{nm})(1 \times 10^{-9}\,\text{m/1 nm})} = 4.74 \times 10^{-19}\,\text{J}$$

6.41 The wavelength of radiation is related to the energy of the electron transition by the equation $\Delta E = hc/\lambda$, where ΔE is the energy difference between the initial orbit with $n = 4$ and the final orbit with $n = 2$. This energy difference is calculated as follows:

$$\Delta E = -hc\,R_H\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right) = (-2.18 \times 10^{-18}\,\text{J})\left(\frac{1}{2^2} - \frac{1}{4^2}\right) = -4.09 \times 10^{-19}\,\text{J}$$

Since the value of ΔE is negative, energy is released. The absolute value of ΔE is used in the calculation of the value of λ:

$$\begin{aligned}\lambda &= \frac{hc}{\Delta E}\\ &= \frac{(6.63 \times 10^{-34}\,\text{J-sec})(3.00 \times 10^{8}\,\text{m/sec})}{4.09 \times 10^{-19}\,\text{J}}\\ &= 4.86 \times 10^{-7}\,\text{m} = 486\,\text{nm}\end{aligned}$$

A wavelength of 4.86×10^{-7} m or 4860 Å belongs in the visible range of light.

6.42 Heisenberg's uncertainty principle negated this postulate. It states that it is impossible to measure the position and velocity of a small particle simultaneously and accurately. In order to describe an electron as having motion in a well-defined circular path, we would have to specify its velocity and position accurately at all times, which, according to the uncertainty principle, is impossible.

6.43 ψ is a mathematical expression that describes the amplitude and the motion of the matter wave of an electron as a function of time. ψ^2 is the probability of finding the electron at some point in space.

6.44 The energy of a hydrogen orbital increases as the value of the principal quantum number increases. Thus, among the pairs, the following have the higher energy: (a) 3*s*; (b) 4*d*; (c) both orbitals, because both have the same principal quantum number. Note: The answer to (c) is true only for the hydrogen atom. For multi-electron atoms the 2*p* orbital is higher in energy than the 2*s*.

6.45 **(a)** 4*d*.
(b) Yes. Any orbital possessing the quantum numbers $n = 4$ and $l = 2$ has the same notation. The $l = 2$ quantum number has associated with it five m_l quantum numbers, 2, 1, 0, −1, −2. Therefore a total of five orbitals have the designation 4*d*.

6.46 For the $n = 4$ principal quantum energy level, orbitals with $l = 0, 1, 2, 3$ values can exist. Thus, the $n = 4$ principal quantum energy level possesses the following subshells: 4*s*, 4*p*, 4*d*, and 4*f*. They all possess the same energy because they have the same principal quantum number, 4.

6.47 For the $n = 2$ principal quantum energy level, orbitals with $l = 0, 1$ values can exist. For each l value, there exists a further subset of orbital types based on their *ml* values:

l	m_l	Shorthand notation
0	0	2*s*
1	+1	2*p*
1	0	2*p*
1	−1	2*p*

6.48 The fact that the surface is completely spherical tells you that it is an *s* type orbital; however, you cannot tell if it is a 1*s* orbital without knowing the total number of spherical nodes. If the electron density is only zero at the nucleus, then it is a 1*s* orbital. If it has another node at a distance from the nucleus it is a 2*s* orbital; if it has two nodes at a distance from the nucleus it is a 3*s* orbital.

6.49 There is sufficient information. The fact that the valence-shell electron configuration includes the 4*p* electrons tells you the element is beyond the transition elements of the fourth period. Only these elements in the fourth period have *p* electrons in the valence shell. Arsenic is the element because it possesses the valence electrons $4s^24p^3$ and it has a set of three 4*p* electrons.

6.50 **(d)** violates the Pauli Principle because a single orbital can possess no more than two electrons.

6.51 A node is a place in space where the amplitude of a wave function is zero. It is generally true that the more nodes a wave function possesses the higher its energy. The size of a hydrogen orbital increases as the value of the principal quantum number increases. A 4*s* orbital is thus larger than a 3*s* orbital. For each orbital with a given *n* principal quantum number, the number of nodes is the same—one less than the value of *n*. Therefore, the 4*s* orbital has three nodes, and the 3*s* orbital has two nodes.

6.52 See Figure 6.24 in the text for these contour representations. These orbitals all have the same energy in a hydrogen atom because they all have the same principal quantum number value, 3.

6.53 The expression $E = hc/\lambda$ is valid for electromagnetic radiation. Matter typically used by humans is not electromagnetic radiation and does not move at the speed of light.

6.54 First, convert megacycles per second into cycles per second using the equivalence 1 megacycle = 1×10^6 cycles

$$v_{\text{microwave}} = \left(20{,}000\,\frac{\text{megacycles}}{\text{sec}}\right) \times \left(\frac{1 \times 10^6\ \text{cycles}}{1\ \text{megacycle}}\right) = 2 \times 10^{10}\ \text{cycles/sec}$$

$$v_{\text{visible}} = \left(100{,}000{,}000\,\frac{\text{megacycles}}{\text{sec}}\right) \times \left(\frac{1 \times 10^6\ \text{cycles}}{1\ \text{megacycle}}\right) = 1 \times 10^{14}\ \text{cycles/sec}$$

Substituting these values into the relation $E = hv$ yields

$$E_{\text{microwave}} = hv = (6.63 \times 10^{-34}\ \text{J-sec})(2 \times 10^{10}/\text{sec}) = 1.33 \times 10^{-23}\ \text{J}$$

$$E_{\text{visible}} = hv = (6.63 \times 10^{-34}\ \text{J-sec})(1 \times 10^{14}/\text{sec}) = 6.63 \times 10^{-20}\ \text{J}$$

The value of the ratio $E_{\text{microwave}}/E_{\text{visible}}$ shows the relative magnitude of their energies:

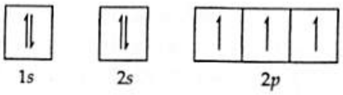
$$\frac{E_{\text{microwave}}}{E_{\text{visible}}} = \frac{1.33 \times 10^{-23}\ \text{J}}{6.63 \times 10^{-20}\ \text{J}} = 2.01 \times 10^{-4}$$

From this we see that $E_{\text{microwave}} = 2.01 \times 10^{-3}\ E_{\text{visible}}$. Thus the energy content of microwave radiation is significantly less than that of visible radiation.

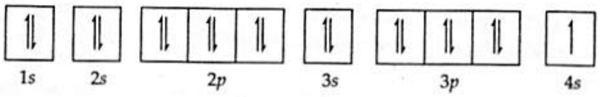
6.55 **(a)** Electrons enter the 2*s* and 2*p* orbitals in the atoms of the second-row elements and the 5*s*, 4*d*, and 5*p* orbitals in the atoms of the fifth period.

(b) The energy of a 3*d* orbital is lower than the energy of a 4*p* orbital. The 3*d* orbital is filled with electrons before the 4*p* orbital when atoms of the fourth period are formed. In order for the fourth row to be composed of elements sequentially increasing by one atomic number, the transition elements with their 3*d* valence electrons must be included.

6.56 **(a)** $1s^2\boxed{2s^22p^3}$—outer orbital set

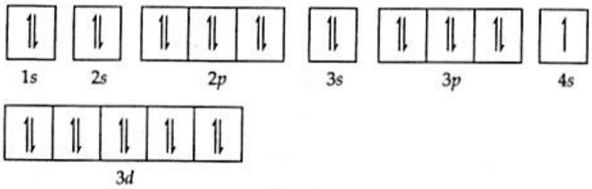


(b) $1s^22s^22p^63s^23p^6\boxed{4s^1}$—outer orbital set

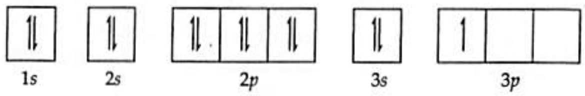


(c) $1s^22s^22p^63s^23p^6\boxed{4s^13d^{10}}$—outer orbital set

The 4*s* and 3*d* orbitals are very similar in energy in the atom once the orbitals are filled with electrons. Thus, both orbitals comprise the outer set of orbitals. The outer orbital electron configuration is $4s^13d^{10}$ instead of the expected $4s^23d^9$ because a half-filled 4*s* orbital and a filled 3*d* orbital appear to be more stable than a partially filled 3*d* orbital.



(d) $1s^22s^22p^6\boxed{3s^23p^1}$—outer orbital set



6.57 **(a)** The element with the highest-energy $4s^24p^2$ electron configuration belongs to the fourth period and the carbon family because $n = 4$ (fourth period) and the valence electron configuration of the carbon family is s^2p^2. The element is germanium (Ge).

(b) The element with the highest-energy $3d^64s^2$ electron configuration belongs to the fourth period and to the transition element family because $n = 4$ for the *s* orbital and transition elements have incompletely filled *d* orbitals. The element is iron ($1s^22s^22p^63s^23p^64s^23d^6$): Fe.

(c) The only family that has an ns^1 electron configuration is the alkali metal family, group 1A. The element belongs to the sixth period because $n = 6$. It is cesium (Cs).

(d) This element belongs to the second period, as $n = 2$ for the valence electrons: Fluorine (F).

(e) See (b) for rationale: Zinc (Zn).

6.58 **(a)** Neon has the electron configuration $1s^22s^22p^6$; argon's is $1s^22s^22p^63s^23p^6$. The electrons that comprise the shell from Ne to Ar are the $3s^23p^6$ electrons.

(b) Li has the electron configuration $1s^22s^1$. The nearest inert gas in the periodic table is He, with a $1s^2$ electron configuration. Li loses the $2s^1$ electron to form Li^+, with a $1s^2$ electron configuration. Sr has the electron configuration [Kr]$5s^2$. Sr loses the two 5*s* electrons to form Sr^{2+} with a [Kr] electron configuration. Oxygen has the electron configuration [He] $2s^22p^4$. The nearest inert gas in the periodic table is Ne. Oxygen adds two 2*p* electrons to form O^{2-} with the neon electron configuration $1s^22s^22p^6$.

6.59 The electron configurations shown in (a) and (c) correspond to excited states. The groundstate electron configuration for (a) is $1s^2$, and for (c), [Ar] $4s^23d^4$. The electron configurations in (b) and (d) correspond to ground states.

6.60 **(a)** This group of elements has six valence electrons and is not a group of transition element. With six valence electrons the group is number 6A[16]. Oxygen is an example with its valence electron configuration of $2s^22p^4$. The other members of the oxygen family will have a similar electron configuration.

(b) This group must contain transition elements because the $(n - 1)d$ orbitals are involved in the outer electron configuration. Choose row four in the periodic table because it has transition elements and count three elements from the left side of the periodic table (the elements in the group possess three outer electrons, 2 + 1 = 3): It is Sc and it has the valence electron configuration $4s^23d^1$. The other members of the group are Y, La, Ac. Lu and Lr also have this outer electron configuration as they are the last elements in the Lanthanides and Actinides, along with a filled set of *nf* orbitals.

(c) Using the approach in (a) and counting two elements from the left in row two you find the element Be. Thus, this group is the alkaline-earth elements. Also He has this electron configuration.

(d) B, Al, Ga, In, and Tl.

6.61 **(a)** Student's (B) conclusion is not reasonable because other elements can also have a valence 4*p* orbital. The value of $n = 4$ for a valence orbital tells us that the element is in the fourth period. The value of $l = 1$ tells us that the electron is a *p* orbital; it is a 4*p* electron. Arsenic has the valence-electron configuration $[\text{Ar}]4s^23d^{10}4p^3$. One of the valence electrons is a 4*p* electron; thus, As is a possibility. However the elements Ga, Ge, Se, Br, and Kr also possess 4*p* valence electrons and could also be the unknown element.

(b) To identify the unknown element student (B) needs either the total number of electrons or the total number of valence electrons possessed by the element.

(c) Yes. The noble or rare-gas elements have a filled valence electronic structure, ns^2np^6, which is very stable. Thus these elements do not readily lose or gain electrons. This information helps identify the unknown element as krypton.

(d) Yes. The only metal in the fourth period with a valence 4*p* electron is gallium.

6.62 **(a)**

$$E = h\frac{c}{\lambda}$$

$$E = \frac{(6.63 \times 10^{-34}\ \text{J-s})\left(3.00 \times 10^{8}\,\frac{\text{m}}{\text{s}}\right)}{(589.3\ \text{nm})\left(\frac{1\ \text{m}}{10^{9}\ \text{nm}}\right)} = 3.38 \times 10^{-19}\ \text{J}$$

(b) The energy calculated in (a) is for a photon emitted by a single sodium atom. To calculate the energy for a mole of sodium atoms you must use Avogadro's number:

$$E = \left(\frac{3.38 \times 10^{-19}\ \text{J}}{\text{Na atom}}\right)\left(\frac{6.022 \times 10^{23}\ \text{Na atoms}}{1\ \text{mole}}\right) = 203{,}000\ \text{J} = 203\ \text{kJ}$$

(c) The electron configuration of sodium is $1s^22s^22p^63s^1$. The valence electron is $3s^1$.

(d) The formation of a positive ion occurs when an electron is removed. To form Na^+ the valence electron is removed: $1s^22s^22p^6$.

6.63 **(a)** Electrons and photons are "screwy" because if we say they behave like particles then this gives the wrong impression; similarly if we say they behave like waves that also gives the wrong impression. They possess a duality of both. This duality behavior on the atomic scale is different and we do not observe it for particles in our macroscopic world.

(b) It is very difficult to find analogies between behaviors of particles in our familiar world and that of electrons and photons at the atomic level; the particles behave entirely different. They do not act in "normal" ways and we find this hard to model. It is likely whatever we visualize or model is not quite correct and the search for truth may never end, leading to our potential torment. Learn how electrons and photons in the atomic world behave; be able to describe their behaviors based on what we know today. Use the models that have been developed and be able to apply them. Recognize the limitations of the models.

6.64 **(c)**

$$v = \frac{c}{\lambda}$$

$$= \frac{3.00 \times 10^{8}\ \text{m/sec}}{(7.52 \times 10^{6}\ \text{Å})(1 \times 10^{-10}\ \text{m/Å})} = 3.99 \times 10^{11}/\text{s}$$

6.65 **(b)**

$$E = hv$$

$$= (6.63 \times 10^{-34}\ \text{J-sec})(9.00 \times 10^{11}/\text{s}) = 5.97 \times 10^{-22}\ \text{J}$$

6.66 **(a)** $\lambda = \dfrac{h}{mv}$

$$= \left(\frac{6.63 \times 10^{-34}\ \text{J-sec}}{(9.1 \times 10^{-28}\ \text{g})(4.00 \times 10^{9}\ \text{cm/sec})}\right) \times \left(\frac{1\ \text{kg-m}^2/\text{sec}^2}{1\ \text{J}}\right)\left(\frac{100\ \text{cm}}{\text{m}}\right)^2\left(\frac{10^{3}\ \text{g}}{\text{kg}}\right) = 1.8 \times 10^{-9}\ \text{cm}$$

The answer can be converted to angstroms as follows:

$$\lambda = (1.8 \times 10^{-9}\ \text{cm})\left(\frac{1\text{Å}}{1 \times 10^{-8}\ \text{cm}}\right) = 0.18\text{Å}$$

6.67 (e). **6.68** (b). **6.69** (e). **6.70** (e). **6.71** (b). **6.72** (a). **6.73** (d).

6.74 (b). The value of m_l cannot be greater than the value of l.

6.75 (d). $\Delta E \propto -\left(\frac{1}{4^2} - \frac{1}{2^2}\right) = \frac{3}{16}$. Energy is absorbed because ΔE is positive.

6.76 (d). **6.77** (b). **6.78** (c).

6.79 (d). In many-electron atoms, the energy of an electron depends on both n and l. The electron with the lowest n value along with the lowest l value has the lowest energy.

6.80 (e). $1s^22s^22p^63s^23p^63d^{10}4s^24p^64d^{10}5s^25p^6$.

6.81 (c).

6.82 (e).

6.83 (e). For $n = 3$, the maximum possible outer electron configuration is $3s^23p^63d^{10}$.

6.84 (c).