

CHAPTER 7

THE ELECTRONIC STRUCTURE OF ATOMS

7.7 (a) $\lambda = \frac{c}{\nu} = \frac{3.00 \times 10^8 \text{ m/s}}{8.6 \times 10^{13} / \text{s}} = 3.5 \times 10^{-6} \text{ m} = 3.5 \times 10^3 \text{ nm}$

(b) $\nu = \frac{c}{\lambda} = \frac{3.00 \times 10^8 \text{ m/s}}{566 \times 10^{-9} \text{ m}} = 5.30 \times 10^{14} / \text{s} = 5.30 \times 10^{14} \text{ Hz}$

7.8 (a)

Strategy: We are given the wavelength of an electromagnetic wave and asked to calculate the frequency. Rearranging Equation (7.1) of the text and replacing u with c (the speed of light) gives:

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

Solution: Because the speed of light is given in meters per second, it is convenient to first convert wavelength to units of meters. Recall that $1 \text{ nm} = 1 \times 10^{-9} \text{ m}$ (see Table 1.3 of the text). We write:

$$456 \text{ nm} \times \frac{1 \times 10^{-9} \text{ m}}{1 \text{ nm}} = 456 \times 10^{-9} \text{ m} = 4.56 \times 10^{-7} \text{ m}$$

Substituting in the wavelength and the speed of light ($3.00 \times 10^8 \text{ m/s}$), the frequency is:

$$\nu = \frac{c}{\lambda} = \frac{3.00 \times 10^8 \frac{\text{m}}{\text{s}}}{4.56 \times 10^{-7} \text{ m}} = 6.58 \times 10^{14} \text{ s}^{-1} \text{ or } 6.58 \times 10^{14} \text{ Hz}$$

Check: The answer shows that 6.58×10^{14} waves pass a fixed point every second. This very high frequency is in accordance with the very high speed of light.

(b)

Strategy: We are given the frequency of an electromagnetic wave and asked to calculate the wavelength. Rearranging Equation (7.1) of the text and replacing u with c (the speed of light) gives:

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

Solution: Substituting in the frequency and the speed of light ($3.00 \times 10^8 \text{ m/s}$) into the above equation, the wavelength is:

$$\lambda = \frac{c}{\nu} = \frac{3.00 \times 10^8 \frac{\text{m}}{\text{s}}}{2.45 \times 10^9 \frac{1}{\text{s}}} = 0.122 \text{ m}$$

The problem asks for the wavelength in units of nanometers. Recall that $1 \text{ nm} = 1 \times 10^{-9} \text{ m}$.

$$\lambda = 0.122 \text{ m} \times \frac{1 \text{ nm}}{1 \times 10^{-9} \text{ m}} = 1.22 \times 10^8 \text{ nm}$$

7.9 Since the speed of light is 3.00×10^8 m/s, we can write

$$(1.3 \times 10^8 \text{ mi}) \times \frac{1.61 \text{ km}}{1 \text{ mi}} \times \frac{1000 \text{ m}}{1 \text{ km}} \times \frac{1 \text{ s}}{3.00 \times 10^8 \text{ m}} = 7.0 \times 10^2 \text{ s}$$

Would the time be different for other types of electromagnetic radiation?

7.10 A radio wave is an electromagnetic wave, which travels at the speed of light. The speed of light is in units of m/s, so let's convert distance from units of miles to meters. ($28 \text{ million mi} = 2.8 \times 10^7 \text{ mi}$)

$$? \text{ distance (m)} = (2.8 \times 10^7 \text{ mi}) \times \frac{1.61 \text{ km}}{1 \text{ mi}} \times \frac{1000 \text{ m}}{1 \text{ km}} = 4.5 \times 10^{10} \text{ m}$$

Now, we can use the speed of light as a conversion factor to convert from meters to seconds ($c = 3.00 \times 10^8$ m/s).

$$? \text{ min} = (4.5 \times 10^{10} \text{ m}) \times \frac{1 \text{ s}}{3.00 \times 10^8 \text{ m}} = 1.5 \times 10^2 \text{ s} = 2.5 \text{ min}$$

$$7.11 \quad \lambda = \frac{c}{\nu} = \frac{3.00 \times 10^8 \text{ m/s}}{9192631770 \text{ s}^{-1}} = 3.26 \times 10^{-2} \text{ m} = 3.26 \times 10^7 \text{ nm}$$

This radiation falls in the microwave region of the spectrum. (See Figure 7.3 of the text.)

7.12 The wavelength is:

$$\lambda = \frac{1 \text{ m}}{1,650,763.73 \text{ wavelengths}} = 6.05780211 \times 10^{-7} \text{ m}$$

$$\nu = \frac{c}{\lambda} = \frac{3.00 \times 10^8 \text{ m/s}}{6.05780211 \times 10^{-7} \text{ m}} = 4.95 \times 10^{14} \text{ s}^{-1}$$

$$7.15 \quad E = h\nu = \frac{hc}{\lambda} = \frac{(6.63 \times 10^{-34} \text{ J}\cdot\text{s})(3.00 \times 10^8 \text{ m/s})}{624 \times 10^{-9} \text{ m}} = 3.19 \times 10^{-19} \text{ J}$$

7.16 (a)

Strategy: We are given the frequency of an electromagnetic wave and asked to calculate the wavelength. Rearranging Equation (7.1) of the text and replacing u with c (the speed of light) gives:

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

Solution: Substituting in the frequency and the speed of light (3.00×10^8 m/s) into the above equation, the wavelength is:

$$\lambda = \frac{3.00 \times 10^8 \frac{\text{m}}{\text{s}}}{7.5 \times 10^{14} \frac{1}{\text{s}}} = 4.0 \times 10^{-7} \text{ m} = 4.0 \times 10^2 \text{ nm}$$

Check: The wavelength of 400 nm calculated is in the blue region of the visible spectrum as expected.

(b)

Strategy: We are given the frequency of an electromagnetic wave and asked to calculate its energy. Equation (7.2) of the text relates the energy and frequency of an electromagnetic wave.

$$E = h\nu$$

Solution: Substituting in the frequency and Planck's constant (6.63×10^{-34} J·s) into the above equation, the energy of a single photon associated with this frequency is:

$$E = h\nu = (6.63 \times 10^{-34} \text{ J}\cdot\text{s}) \left(7.5 \times 10^{14} \frac{1}{\text{s}} \right) = \mathbf{5.0 \times 10^{-19} \text{ J}}$$

Check: We expect the energy of a single photon to be a very small energy as calculated above, 5.0×10^{-19} J.

$$7.17 \quad \text{(a)} \quad \lambda = \frac{c}{\nu} = \frac{3.00 \times 10^8 \text{ m/s}}{6.0 \times 10^4 /\text{s}} = \mathbf{5.0 \times 10^3 \text{ m} = 5.0 \times 10^{12} \text{ nm}}$$

The radiation does not fall in the visible region; it is radio radiation. (See Figure 7.3 of the text.)

$$\text{(b)} \quad E = h\nu = (6.63 \times 10^{-34} \text{ J}\cdot\text{s})(6.0 \times 10^4 /\text{s}) = \mathbf{4.0 \times 10^{-29} \text{ J}}$$

$$\text{(c)} \quad \text{Converting to J/mol: } E = \frac{4.0 \times 10^{-29} \text{ J}}{1 \text{ photon}} \times \frac{6.022 \times 10^{23} \text{ photons}}{1 \text{ mol}} = \mathbf{2.4 \times 10^{-5} \text{ J/mol}}$$

7.18 The energy given in this problem is for 1 mole of photons. To apply $E = h\nu$, we must divide the energy by Avogadro's number. The energy of one photon is:

$$E = \frac{1.0 \times 10^3 \text{ kJ}}{1 \text{ mol}} \times \frac{1 \text{ mol}}{6.022 \times 10^{23} \text{ photons}} \times \frac{1000 \text{ J}}{1 \text{ kJ}} = 1.7 \times 10^{-18} \text{ J/photon}$$

The wavelength of this photon can be found using the relationship, $E = \frac{hc}{\lambda}$.

$$\lambda = \frac{hc}{E} = \frac{(6.63 \times 10^{-34} \text{ J}\cdot\text{s}) \left(3.00 \times 10^8 \frac{\text{m}}{\text{s}} \right)}{1.7 \times 10^{-18} \text{ J}} = 1.2 \times 10^{-7} \text{ m} \times \frac{1 \text{ nm}}{1 \times 10^{-9} \text{ m}} = \mathbf{1.2 \times 10^2 \text{ nm}}$$

The radiation is in the **ultraviolet** region (see Figure 7.3 of the text).

$$7.19 \quad E = h\nu = \frac{hc}{\lambda} = \frac{(6.63 \times 10^{-34} \text{ J}\cdot\text{s})(3.00 \times 10^8 \text{ m/s})}{(0.154 \times 10^{-9} \text{ m})} = \mathbf{1.29 \times 10^{-15} \text{ J}}$$

$$7.20 \quad \text{(a)} \quad \lambda = \frac{c}{\nu}$$

$$\lambda = \frac{3.00 \times 10^8 \frac{\text{m}}{\text{s}}}{8.11 \times 10^{14} \frac{1}{\text{s}}} = 3.70 \times 10^{-7} \text{ m} = \mathbf{3.70 \times 10^2 \text{ nm}}$$

(b) Checking Figure 7.3 of the text, you should find that the visible region of the spectrum runs from 400 to 700 nm. 370 nm is in the **ultraviolet** region of the spectrum.

- (c) $E = h\nu$. Substitute the frequency (ν) into this equation to solve for the energy of one quantum associated with this frequency.

$$E = h\nu = (6.63 \times 10^{-34} \text{ J}\cdot\text{s}) \left(8.11 \times 10^{14} \frac{1}{\text{s}} \right) = 5.38 \times 10^{-19} \text{ J}$$

7.25 The arrangement of energy levels for each element is unique. The frequencies of light emitted by an element are characteristic of that element. Even the frequencies emitted by isotopes of the same element are very slightly different.

7.26 The emitted light could be analyzed by passing it through a prism.

7.27 Light emitted by fluorescent materials always has lower energy than the light striking the fluorescent substance. Absorption of visible light could not give rise to emitted ultraviolet light because the latter has higher energy.

The reverse process, ultraviolet light producing visible light by fluorescence, is very common. Certain brands of laundry detergents contain materials called “optical brighteners” which, for example, can make a white shirt look much whiter and brighter than a similar shirt washed in ordinary detergent.

7.28 Excited atoms of the chemical elements emit the same characteristic frequencies or lines in a terrestrial laboratory, in the sun, or in a star many light-years distant from earth.

7.29 (a) The energy difference between states E_1 and E_4 is:

$$E_4 - E_1 = (-1.0 \times 10^{-19} \text{ J}) - (-15 \times 10^{-19} \text{ J}) = 14 \times 10^{-19} \text{ J}$$

$$\lambda = \frac{hc}{\Delta E} = \frac{(6.63 \times 10^{-34} \text{ J}\cdot\text{s})(3.00 \times 10^8 \text{ m/s})}{14 \times 10^{-19} \text{ J}} = 1.4 \times 10^{-7} \text{ m} = 1.4 \times 10^2 \text{ nm}$$

(b) The energy difference between the states E_2 and E_3 is:

$$E_3 - E_2 = (-5.0 \times 10^{-19} \text{ J}) - (-10.0 \times 10^{-19} \text{ J}) = 5 \times 10^{-19} \text{ J}$$

(c) The energy difference between the states E_1 and E_3 is:

$$E_1 - E_3 = (-15 \times 10^{-19} \text{ J}) - (-5.0 \times 10^{-19} \text{ J}) = -10 \times 10^{-19} \text{ J}$$

Ignoring the negative sign of ΔE , the wavelength is found as in part (a).

$$\lambda = \frac{hc}{\Delta E} = \frac{(6.63 \times 10^{-34} \text{ J}\cdot\text{s})(3.00 \times 10^8 \text{ m/s})}{10 \times 10^{-19} \text{ J}} = 2.0 \times 10^{-7} \text{ m} = 2.0 \times 10^2 \text{ nm}$$

7.30 We use more accurate values of h and c for this problem.

$$E = \frac{hc}{\lambda} = \frac{(6.6256 \times 10^{-34} \text{ J}\cdot\text{s})(2.998 \times 10^8 \text{ m/s})}{656.3 \times 10^{-9} \text{ m}} = 3.027 \times 10^{-19} \text{ J}$$

7.31 In this problem $n_i = 5$ and $n_f = 3$.

$$\Delta E = R_H \left(\frac{1}{n_i^2} - \frac{1}{n_f^2} \right) = (2.18 \times 10^{-18} \text{ J}) \left(\frac{1}{5^2} - \frac{1}{3^2} \right) = -1.55 \times 10^{-19} \text{ J}$$

The sign of ΔE means that this is energy associated with an emission process.

$$\lambda = \frac{hc}{\Delta E} = \frac{(6.63 \times 10^{-34} \text{ J}\cdot\text{s})(3.00 \times 10^8 \text{ m/s})}{1.55 \times 10^{-19} \text{ J}} = 1.28 \times 10^{-6} \text{ m} = 1.28 \times 10^3 \text{ nm}$$

Is the sign of the energy change consistent with the sign conventions for exo- and endothermic processes?

- 7.32 Strategy:** We are given the initial and final states in the emission process. We can calculate the energy of the emitted photon using Equation (7.5) of the text. Then, from this energy, we can solve for the frequency of the photon, and from the frequency we can solve for the wavelength. The value of Rydberg's constant is $2.18 \times 10^{-18} \text{ J}$.

Solution: From Equation (7.5) we write:

$$\begin{aligned}\Delta E &= R_{\text{H}} \left(\frac{1}{n_i^2} - \frac{1}{n_f^2} \right) \\ \Delta E &= (2.18 \times 10^{-18} \text{ J}) \left(\frac{1}{4^2} - \frac{1}{2^2} \right) \\ \Delta E &= -4.09 \times 10^{-19} \text{ J}\end{aligned}$$

The negative sign for ΔE indicates that this is energy associated with an emission process. To calculate the frequency, we will omit the minus sign for ΔE because the frequency of the photon must be positive. We know that

$$\Delta E = h\nu$$

Rearranging the equation and substituting in the known values,

$$\nu = \frac{\Delta E}{h} = \frac{(4.09 \times 10^{-19} \text{ J})}{(6.63 \times 10^{-34} \text{ J}\cdot\text{s})} = 6.17 \times 10^{14} \text{ s}^{-1} \text{ or } 6.17 \times 10^{14} \text{ Hz}$$

We also know that $\lambda = \frac{c}{\nu}$. Substituting the frequency calculated above into this equation gives:

$$\lambda = \frac{3.00 \times 10^8 \frac{\text{m}}{\text{s}}}{\left(6.17 \times 10^{14} \frac{1}{\text{s}} \right)} = 4.86 \times 10^{-7} \text{ m} = 486 \text{ nm}$$

Check: This wavelength is in the visible region of the electromagnetic region (see Figure 7.3 of the text). This is consistent with the fact that because $n_i = 4$ and $n_f = 2$, this transition gives rise to a spectral line in the Balmer series (see Figure 7.9 of the text).

- 7.33** This problem must be worked to four significant figure accuracy. We use $6.6256 \times 10^{-34} \text{ J}\cdot\text{s}$ for Planck's constant and $2.998 \times 10^8 \text{ m/s}$ for the speed of light. First calculate the energy of each of the photons.

$$\begin{aligned}E &= \frac{hc}{\lambda} = \frac{(6.6256 \times 10^{-34} \text{ J}\cdot\text{s})(2.998 \times 10^8 \text{ m/s})}{589.0 \times 10^{-9} \text{ m}} = 3.372 \times 10^{-19} \text{ J} \\ E &= \frac{hc}{\lambda} = \frac{(6.6256 \times 10^{-34} \text{ J}\cdot\text{s})(2.998 \times 10^8 \text{ m/s})}{589.6 \times 10^{-9} \text{ m}} = 3.369 \times 10^{-19} \text{ J}\end{aligned}$$

For *one* photon the energy difference is:

$$\Delta E = (3.372 \times 10^{-19} \text{ J}) - (3.369 \times 10^{-19} \text{ J}) = 3 \times 10^{-22} \text{ J}$$

For *one mole* of photons the energy difference is:

$$\frac{3 \times 10^{-22} \text{ J}}{1 \text{ photon}} \times \frac{6.022 \times 10^{23} \text{ photons}}{1 \text{ mol}} = 2 \times 10^2 \text{ J/mol}$$

$$7.34 \quad \Delta E = R_H \left(\frac{1}{n_i^2} - \frac{1}{n_f^2} \right)$$

n_f is given in the problem and R_H is a constant, but we need to calculate ΔE . The photon energy is:

$$E = \frac{hc}{\lambda} = \frac{(6.63 \times 10^{-34} \text{ J}\cdot\text{s})(3.00 \times 10^8 \text{ m/s})}{434 \times 10^{-9} \text{ m}} = 4.58 \times 10^{-19} \text{ J}$$

Since this is an emission process, the energy change ΔE must be negative, or $-4.58 \times 10^{-19} \text{ J}$.

Substitute ΔE into the following equation, and solve for n_i .

$$\begin{aligned} \Delta E &= R_H \left(\frac{1}{n_i^2} - \frac{1}{n_f^2} \right) \\ -4.58 \times 10^{-19} \text{ J} &= (2.18 \times 10^{-18} \text{ J}) \left(\frac{1}{n_i^2} - \frac{1}{2^2} \right) \\ \frac{1}{n_i^2} &= \left(\frac{-4.58 \times 10^{-19} \text{ J}}{2.18 \times 10^{-18} \text{ J}} \right) + \frac{1}{2^2} = -0.210 + 0.250 = 0.040 \\ n_i &= \frac{1}{\sqrt{0.040}} = 5 \end{aligned}$$

$$7.39 \quad \lambda = \frac{h}{mu} = \frac{6.63 \times 10^{-34} \text{ J}\cdot\text{s}}{(1.675 \times 10^{-27} \text{ kg})(7.00 \times 10^2 \text{ m/s})} = 5.65 \times 10^{-10} \text{ m} = \mathbf{0.565 \text{ nm}}$$

7.40 Strategy: We are given the mass and the speed of the proton and asked to calculate the wavelength. We need the de Broglie equation, which is Equation (7.7) of the text. Note that because the units of Planck's constant are J·s, m must be in kg and u must be in m/s (1 J = 1 kg·m²/s²).

Solution: Using Equation (7.7) we write:

$$\begin{aligned} \lambda &= \frac{h}{mu} \\ \lambda &= \frac{h}{mu} = \frac{\left(6.63 \times 10^{-34} \frac{\text{kg}\cdot\text{m}^2}{\text{s}^2} \cdot \text{s} \right)}{(1.673 \times 10^{-27} \text{ kg})(2.90 \times 10^8 \text{ m/s})} = 1.37 \times 10^{-15} \text{ m} \end{aligned}$$

The problem asks to express the wavelength in nanometers.

$$\lambda = (1.37 \times 10^{-15} \text{ m}) \times \frac{1 \text{ nm}}{1 \times 10^{-9} \text{ m}} = \mathbf{1.37 \times 10^{-6} \text{ nm}}$$

7.41 Converting the velocity to units of m/s::

$$\frac{1.20 \times 10^2 \text{ mi}}{1 \text{ hr}} \times \frac{1.61 \text{ km}}{1 \text{ mi}} \times \frac{1000 \text{ m}}{1 \text{ km}} \times \frac{1 \text{ hr}}{3600 \text{ s}} = 53.7 \text{ m/s}$$

$$\lambda = \frac{h}{mu} = \frac{6.63 \times 10^{-34} \text{ J}\cdot\text{s}}{(0.0124 \text{ kg})(53.7 \text{ m/s})} = 9.96 \times 10^{-34} \text{ m} = \mathbf{9.96 \times 10^{-32} \text{ cm}}$$

7.42 First, we convert mph to m/s.

$$\frac{35 \text{ mi}}{1 \text{ h}} \times \frac{1.61 \text{ km}}{1 \text{ mi}} \times \frac{1000 \text{ m}}{1 \text{ km}} \times \frac{1 \text{ h}}{3600 \text{ s}} = 16 \text{ m/s}$$

$$\lambda = \frac{h}{mu} = \frac{\left(6.63 \times 10^{-34} \frac{\text{kg}\cdot\text{m}^2}{\text{s}^2}\cdot\text{s}\right)}{(2.5 \times 10^{-3} \text{ kg})(16 \text{ m/s})} = 1.7 \times 10^{-32} \text{ m} = \mathbf{1.7 \times 10^{-23} \text{ nm}}$$

7.53 The angular momentum quantum number l can have integral (i.e. whole number) values from 0 to $n - 1$. In this case $n = 2$, so the allowed values of the angular momentum quantum number, l , are **0** and **1**.

Each allowed value of the angular momentum quantum number labels a subshell. Within a given subshell (label l) there are $2l + 1$ allowed energy states (orbitals) each labeled by a different value of the magnetic quantum number. The allowed values run from $-l$ through 0 to $+l$ (whole numbers only). For the subshell labeled by the angular momentum quantum number $l = 1$, the allowed values of the magnetic quantum number, m_l , are **-1, 0, and 1**. For the other subshell in this problem labeled by the angular momentum quantum number $l = 0$, the allowed value of the magnetic quantum number is **0**.

If the allowed whole number values run from -1 to $+1$, are there always $2l + 1$ values? Why?

7.54 Strategy: What are the relationships among n , l , and m_l ?

Solution: We are given the principal quantum number, $n = 3$. The possible l values range from 0 to $(n - 1)$. Thus, there are three possible values of l : 0, 1, and 2, corresponding to the s , p , and d orbitals, respectively. The values of m_l can vary from $-l$ to l . The values of m_l for each l value are:

$$l = 0: \quad m_l = 0 \qquad l = 1: \quad m_l = -1, 0, 1 \qquad l = 2: \quad m_l = -2, -1, 0, 1, 2$$

- 7.55** (a) $2p$: $n = 2$, $l = 1$, $m_l = 1, 0$, or -1
 (b) $3s$: $n = 3$, $l = 0$, $m_l = 0$ (only allowed value)
 (c) $5d$: $n = 5$, $l = 2$, $m_l = 2, 1, 0, -1$, or -2

An orbital in a subshell can have any of the allowed values of the magnetic quantum number for that subshell. All the orbitals in a subshell have exactly the same energy.

- 7.56** (a) The number given in the designation of the subshell is the principal quantum number, so in this case $n = 3$. For s orbitals, $l = 0$. m_l can have integer values from $-l$ to $+l$, therefore, $m_l = 0$. The electron spin quantum number, m_s , can be either $+1/2$ or $-1/2$.

Following the same reasoning as part (a)

(b) $4p$: $n = 4$; $l = 1$; $m_l = -1, 0, 1$; $m_s = +1/2, -1/2$

(c) $3d$: $n = 3$; $l = 2$; $m_l = -2, -1, 0, 1, 2$; $m_s = +1/2, -1/2$

7.57 A $2s$ orbital is larger than a $1s$ orbital. Both have the same spherical shape. The $1s$ orbital is lower in energy than the $2s$.

7.58 The two orbitals are identical in size, shape, and energy. They differ only in their orientation with respect to each other.

Can you assign a specific value of the magnetic quantum number to these orbitals? What are the allowed values of the magnetic quantum number for the $2p$ subshell?

7.59 The allowed values of l are 0, 1, 2, 3, and 4. These correspond to the $5s$, $5p$, $5d$, $5f$, and $5g$ subshells. These subshells each have one, three, five, seven, and nine orbitals, respectively.

7.60 For $n = 6$, the allowed values of l are 0, 1, 2, 3, 4, and 5 [$l = 0$ to $(n - 1)$, integer values]. These l values correspond to the $6s$, $6p$, $6d$, $6f$, $6g$, and $6h$ subshells. These subshells each have 1, 3, 5, 7, 9, and 11 orbitals, respectively (number of orbitals = $2l + 1$).

7.61 There can be a maximum of two electrons occupying one orbital.

(a) two (b) six (c) ten (d) fourteen

What rule of nature demands a maximum of two electrons per orbital? Do they have the same energy? How are they different? Would five $4d$ orbitals hold as many electrons as five $3d$ orbitals? In other words, does the principal quantum number n affect the number of electrons in a given subshell?

<u>n value</u>	<u>orbital sum</u>	<u>total number of electrons</u>
1	1	2
2	$1 + 3 = 4$	8
3	$1 + 3 + 5 = 9$	18
4	$1 + 3 + 5 + 7 = 16$	32
5	$1 + 3 + 5 + 7 + 9 = 25$	50
6	$1 + 3 + 5 + 7 + 9 + 11 = 36$	72

In each case the total number of orbitals is just the square of the n value (n^2). The total number of electrons is $2n^2$.

7.63 $3s$: two $3d$: ten $4p$: six $4f$: fourteen $5f$: fourteen

7.64 The electron configurations for the elements are

(a) N: $1s^2 2s^2 2p^3$ There are three p -type electrons.

(b) Si: $1s^2 2s^2 2p^6 3s^2 3p^2$ There are six s -type electrons.

(c) S: $1s^2 2s^2 2p^6 3s^2 3p^4$ There are no d -type electrons.

7.65 See Figure 7.19 in your textbook.

7.66 In the many-electron atom, the $3p$ orbital electrons are more effectively shielded by the inner electrons of the atom (that is, the $1s$, $2s$, and $2p$ electrons) than the $3s$ electrons. The $3s$ orbital is said to be more “penetrating” than the $3p$ and $3d$ orbitals. In the hydrogen atom there is only one electron, so the $3s$, $3p$, and $3d$ orbitals have the same energy.

7.67 Equation (7.4) of the text gives the orbital energy in terms of the principal quantum number, n , alone (for the hydrogen atom). The energy does not depend on any of the other quantum numbers. If two orbitals in the hydrogen atom have the same value of n , they have equal energy.

(a) $2s > 1s$ (b) $3p > 2p$ (c) equal (d) equal (e) $5s > 4f$

7.68 (a) $2s < 2p$ (b) $3p < 3d$ (c) $3s < 4s$ (d) $4d < 5f$

7.79 (a) is wrong because the magnetic quantum number m_l can have only whole number values.

(c) is wrong because the maximum value of the angular momentum quantum number l is $n - 1$.

(e) is wrong because the electron spin quantum number m_s can have only half-integral values.

7.80 For aluminum, there are not enough electrons in the $2p$ subshell. (The $2p$ subshell holds six electrons.) The number of electrons (13) is correct. The electron configuration should be $1s^2 2s^2 2p^6 3s^2 3p^1$. The configuration shown might be an excited state of an aluminum atom.

For boron, there are too many electrons. (Boron only has five electrons.) The electron configuration should be $1s^2 2s^2 2p^1$. What would be the electric charge of a boron ion with the electron arrangement given in the problem?

For fluorine, there are also too many electrons. (Fluorine only has nine electrons.) The configuration shown is that of the F^- ion. The correct electron configuration is $1s^2 2s^2 2p^5$.

7.81 Since the atomic number is odd, it is mathematically impossible for all the electrons to be paired. There must be at least one that is unpaired. The element would be paramagnetic.

7.82 You should write the electron configurations for each of these elements to answer this question. In some cases, an orbital diagram may be helpful.

B: $[\text{He}]2s^2 2p^1$ (1 unpaired electron)

Ne: (0 unpaired electrons, Why?)

P: $[\text{Ne}]3s^2 3p^3$ (3 unpaired electrons)

Sc: $[\text{Ar}]4s^2 3d^1$ (1 unpaired electron)

Mn: $[\text{Ar}]4s^2 3d^5$ (5 unpaired electrons)

Se: $[\text{Ar}]4s^2 3d^{10} 4p^4$ (2 unpaired electrons)

Kr: (0 unpaired electrons)

Fe: $[\text{Ar}]4s^2 3d^6$ (4 unpaired electrons)

Cd: $[\text{Kr}]5s^2 4d^{10}$ (0 unpaired electrons)

I: $[\text{Kr}]5s^2 4d^{10} 5p^5$ (1 unpaired electron)

Pb: $[\text{Xe}]6s^2 4f^{14} 5d^{10} 6p^2$ (2 unpaired electrons)

7.83 B: $1s^2 2s^2 2p^1$ As: $[\text{Ar}]4s^2 3d^{10} 4p^3$

V: $[\text{Ar}]4s^2 3d^3$ I: $[\text{Kr}]5s^2 4d^{10} 5p^5$

Ni: $[\text{Ar}]4s^2 3d^8$ Au: $[\text{Xe}]6s^1 4f^{14} 5d^{10}$

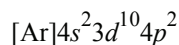
What is the meaning of “[Ar]”? of “[Kr]”? of “[Xe]”?

7.84 Strategy: How many electrons are in the Ge atom ($Z = 32$)? We start with $n = 1$ and proceed to fill orbitals in the order shown in Figure 7.20 of the text. Remember that any given orbital can hold at most 2 electrons. However, don't forget about degenerate orbitals. Starting with $n = 2$, there are three p orbitals of equal energy, corresponding to $m_l = -1, 0, 1$. Starting with $n = 3$, there are five d orbitals of equal energy, corresponding to $m_l = -2, -1, 0, 1, 2$. We can place electrons in the orbitals according to the Pauli exclusion principle and Hund's rule. The task is simplified if we use the noble gas core preceding Ge for the inner electrons.

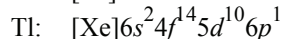
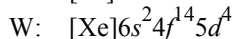
Solution: Germanium has 32 electrons. The noble gas core in this case is [Ar]. (Ar is the noble gas in the period preceding germanium.) [Ar] represents $1s^2 2s^2 2p^6 3s^2 3p^6$. This core accounts for 18 electrons, which leaves 14 electrons to place.

See Figure 7.20 of your text to check the order of filling subshells past the Ar noble gas core. You should find that the order of filling is $4s, 3d, 4p$. There are 14 remaining electrons to distribute among these orbitals. The $4s$ orbital can hold two electrons. Each of the five $3d$ orbitals can hold two electrons for a total of 10 electrons. This leaves two electrons to place in the $4p$ orbitals.

The electrons configuration for Ge is:



You should follow the same reasoning for the remaining atoms.

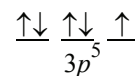
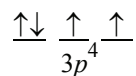
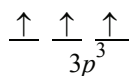


7.85 There are a total of twelve electrons:

Orbital	n	l	m_l	m_s
1s	1	0	0	$+\frac{1}{2}$
1s	1	0	0	$-\frac{1}{2}$
2s	2	0	0	$+\frac{1}{2}$
2s	2	0	0	$-\frac{1}{2}$
2p	2	1	1	$+\frac{1}{2}$
2p	2	1	1	$-\frac{1}{2}$
2p	2	1	0	$+\frac{1}{2}$
2p	2	1	0	$-\frac{1}{2}$
2p	2	1	-1	$+\frac{1}{2}$
2p	2	1	-1	$-\frac{1}{2}$
3s	3	0	0	$+\frac{1}{2}$
3s	3	0	0	$-\frac{1}{2}$

The element is magnesium.

7.86

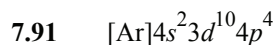


S^+ (5 valence electrons)
3 unpaired electrons

S (6 valence electrons)
2 unpaired electrons

S^- (7 valence electrons)
1 unpaired electron

S^+ has the most unpaired electrons



7.92 The ground state electron configuration of Tc is: $[\text{Kr}]5s^24d^5$.

7.93 We first calculate the wavelength, then we find the color using Figure 7.3 of the text.

$$\lambda = \frac{hc}{E} = \frac{(6.63 \times 10^{-34} \text{ J}\cdot\text{s})(3.00 \times 10^8 \text{ m/s})}{4.30 \times 10^{-19} \text{ J}} = 4.63 \times 10^{-7} \text{ m} = 463 \text{ nm, which is blue.}$$

- 7.94 (a) With $n = 2$, there are n^2 orbitals $= 2^2 = 4$. $m_s = +1/2$, specifies 1 electron per orbital, for a total of **4 electrons**.
- (b) $n = 4$ and $m_l = +1$, specifies one orbital in each subshell with $l = 1, 2$, or 3 (i.e., a $4p$, $4d$, and $4f$ orbital). Each of the three orbitals holds 2 electrons for a total of **6 electrons**.
- (c) If $n = 3$ and $l = 2$, m_l has the values 2, 1, 0, -1, or -2. Each of the five orbitals can hold 2 electrons for a total of **10 electrons** ($2 e^-$ in each of the five $3d$ orbitals).
- (d) If $n = 2$ and $l = 0$, then m_l can only be zero. $m_s = -1/2$ specifies 1 electron in this orbital for a total of **1 electron** (one e^- in the $2s$ orbital).
- (e) $n = 4$, $l = 3$ and $m_l = -2$, specifies one $4f$ orbital. This orbital can hold **2 electrons**.

7.95 See the appropriate sections of the textbook in Chapter 7.

7.96 The wave properties of electrons are used in the operation of an electron microscope.

7.97 It takes:

$$(5.0 \times 10^2 \text{ g ice}) \times \frac{334 \text{ J}}{1 \text{ g ice}} = 1.67 \times 10^5 \text{ J to melt } 5.0 \times 10^2 \text{ g of ice.}$$

Energy of a photon with a wavelength of 660 nm:

$$E = \frac{hc}{\lambda} = \frac{(6.63 \times 10^{-34} \text{ J}\cdot\text{s})(3.00 \times 10^8 \text{ m/s})}{660 \times 10^{-9} \text{ m}} = 3.01 \times 10^{-19} \text{ J}$$

Number of photons needed to melt 5.0×10^2 g of ice:

$$(1.67 \times 10^5 \text{ J}) \times \frac{1 \text{ photon}}{3.01 \times 10^{-19} \text{ J}} = \mathbf{5.5 \times 10^{23} \text{ photons}}$$

The number of water molecules is:

$$(5.0 \times 10^2 \text{ g H}_2\text{O}) \times \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \times \frac{6.022 \times 10^{23} \text{ H}_2\text{O molecules}}{1 \text{ mol H}_2\text{O}} = 1.7 \times 10^{25} \text{ H}_2\text{O molecules}$$

The number of water molecules converted from ice to water by one photon is:

$$\frac{1.7 \times 10^{25} \text{ H}_2\text{O molecules}}{5.5 \times 10^{23} \text{ photons}} = \mathbf{31 \text{ H}_2\text{O molecules/photon}}$$

- 7.98 (a)** First convert 100 mph to units of m/s.

$$\frac{100 \text{ mi}}{1 \text{ h}} \times \frac{1 \text{ h}}{3600 \text{ s}} \times \frac{1.609 \text{ km}}{1 \text{ mi}} \times \frac{1000 \text{ m}}{1 \text{ km}} = 44.7 \text{ m/s}$$

Using the de Broglie equation:

$$\lambda = \frac{h}{mu} = \frac{\left(6.63 \times 10^{-34} \frac{\text{kg} \cdot \text{m}^2}{\text{s}^2} \cdot \text{s}\right)}{(0.141 \text{ kg})(44.7 \text{ m/s})} = 1.05 \times 10^{-34} \text{ m} = \mathbf{1.05 \times 10^{-25} \text{ nm}}$$

- (b)** The average mass of a hydrogen atom is:

$$\frac{1.008 \text{ g}}{1 \text{ mol}} \times \frac{1 \text{ mol}}{6.022 \times 10^{23} \text{ atoms}} = 1.674 \times 10^{-24} \text{ g/H atom} = 1.674 \times 10^{-27} \text{ kg}$$

$$\lambda = \frac{h}{mu} = \frac{\left(6.63 \times 10^{-34} \frac{\text{kg} \cdot \text{m}^2}{\text{s}^2} \cdot \text{s}\right)}{(1.674 \times 10^{-27} \text{ kg})(44.7 \text{ m/s})} = 8.86 \times 10^{-9} \text{ m} = \mathbf{8.86 \text{ nm}}$$

- 7.99** There are many more paramagnetic elements than diamagnetic elements because of Hund's rule.

- 7.100 (a)** First, we can calculate the energy of a single photon with a wavelength of 633 nm.

$$E = \frac{hc}{\lambda} = \frac{(6.63 \times 10^{-34} \text{ J} \cdot \text{s})(3.00 \times 10^8 \text{ m/s})}{633 \times 10^{-9} \text{ m}} = 3.14 \times 10^{-19} \text{ J}$$

The number of photons produced in a 0.376 J pulse is:

$$0.376 \text{ J} \times \frac{1 \text{ photon}}{3.14 \times 10^{-19} \text{ J}} = \mathbf{1.20 \times 10^{18} \text{ photons}}$$

- (b)** Since a 1 W = 1 J/s, the power delivered per a 1.00×10^{-9} s pulse is:

$$\frac{0.376 \text{ J}}{1.00 \times 10^{-9} \text{ s}} = 3.76 \times 10^8 \text{ J/s} = \mathbf{3.76 \times 10^8 \text{ W}}$$

Compare this with the power delivered by a 100-W light bulb!

- 7.101** The energy required to heat the water is: $ms\Delta t = (368 \text{ g})(4.184 \text{ J/g} \cdot ^\circ\text{C})(5.00^\circ\text{C}) = 7.70 \times 10^3 \text{ J}$

Energy of a photon with a wavelength = $1.06 \times 10^4 \text{ nm}$:

$$E = h\nu = \frac{hc}{\lambda} = \frac{(6.63 \times 10^{-34} \text{ J} \cdot \text{s})(3.00 \times 10^8 \text{ m/s})}{1.06 \times 10^{-5} \text{ m}} = 1.88 \times 10^{-20} \text{ J/photon}$$

The number of photons required is:

$$(7.70 \times 10^3 \text{ J}) \times \frac{1 \text{ photon}}{1.88 \times 10^{-20} \text{ J}} = \mathbf{4.10 \times 10^{23} \text{ photons}}$$

7.102 First, let's find the energy needed to photodissociate one water molecule.

$$\frac{285.8 \text{ kJ}}{1 \text{ mol}} \times \frac{1 \text{ mol}}{6.022 \times 10^{23} \text{ molecules}} = 4.746 \times 10^{-22} \text{ kJ/molecule} = 4.746 \times 10^{-19} \text{ J/molecule}$$

The maximum wavelength of a photon that would provide the above energy is:

$$\lambda = \frac{hc}{E} = \frac{(6.63 \times 10^{-34} \text{ J}\cdot\text{s})(3.00 \times 10^8 \text{ m/s})}{4.746 \times 10^{-19} \text{ J}} = \mathbf{4.19 \times 10^{-7} \text{ m} = 419 \text{ nm}}$$

This wavelength is in the visible region of the electromagnetic spectrum. Since water is continuously being struck by visible radiation *without* decomposition, it seems unlikely that photodissociation of water by this method is feasible.

7.103 For the Lyman series, we want the longest wavelength (smallest energy), with $n_i = 2$ and $n_f = 1$. Using Equation (7.5) of the text:

$$\Delta E = R_H \left(\frac{1}{n_i^2} - \frac{1}{n_f^2} \right) = (2.18 \times 10^{-18} \text{ J}) \left(\frac{1}{2^2} - \frac{1}{1^2} \right) = -1.64 \times 10^{-18} \text{ J}$$

$$\lambda = \frac{hc}{\Delta E} = \frac{(6.63 \times 10^{-34} \text{ J}\cdot\text{s})(3.00 \times 10^8 \text{ m/s})}{1.64 \times 10^{-18} \text{ J}} = \mathbf{1.21 \times 10^{-7} \text{ m} = 121 \text{ nm}}$$

For the Balmer series, we want the shortest wavelength (highest energy), with $n_i = \infty$ and $n_f = 2$.

$$\Delta E = R_H \left(\frac{1}{n_i^2} - \frac{1}{n_f^2} \right) = (2.18 \times 10^{-18} \text{ J}) \left(\frac{1}{\infty^2} - \frac{1}{2^2} \right) = -5.45 \times 10^{-19} \text{ J}$$

$$\lambda = \frac{hc}{\Delta E} = \frac{(6.63 \times 10^{-34} \text{ J}\cdot\text{s})(3.00 \times 10^8 \text{ m/s})}{5.45 \times 10^{-19} \text{ J}} = \mathbf{3.65 \times 10^{-7} \text{ m} = 365 \text{ nm}}$$

Therefore the two series do not overlap.

7.104 Since $1 \text{ W} = 1 \text{ J/s}$, the energy output of the light bulb in 1 second is 75 J. The actual energy converted to visible light is 15 percent of this value or 11 J.

First, we need to calculate the energy of one 550 nm photon. Then, we can determine how many photons are needed to provide 11 J of energy.

The energy of one 550 nm photon is:

$$E = \frac{hc}{\lambda} = \frac{(6.63 \times 10^{-34} \text{ J}\cdot\text{s})(3.00 \times 10^8 \text{ m/s})}{550 \times 10^{-9} \text{ m}} = 3.62 \times 10^{-19} \text{ J/photon}$$

The number of photons needed to produce 11 J of energy is:

$$11 \text{ J} \times \frac{1 \text{ photon}}{3.62 \times 10^{-19} \text{ J}} = \mathbf{3.0 \times 10^{19} \text{ photons}}$$

7.105 The heat needed to raise the temperature of 150 mL of water from 20°C to 100°C is:

$$q = ms\Delta t = (150 \text{ g})(4.184 \text{ J/g}\cdot\text{°C})(100 - 20)\text{°C} = 5.0 \times 10^4 \text{ J}$$

The microwave will need to supply more energy than this because only 92.0% of microwave energy is converted to thermal energy of water. The energy that needs to be supplied by the microwave is:

$$\frac{5.0 \times 10^4 \text{ J}}{0.920} = 5.4 \times 10^4 \text{ J}$$

The energy supplied by one photon with a wavelength of $1.22 \times 10^8 \text{ nm}$ (0.122 m) is:

$$E = \frac{hc}{\lambda} = \frac{(6.63 \times 10^{-34} \text{ J}\cdot\text{s})(3.00 \times 10^8 \text{ m/s})}{(0.122 \text{ m})} = 1.63 \times 10^{-24} \text{ J}$$

The number of photons needed to supply $5.4 \times 10^4 \text{ J}$ of energy is:

$$(5.4 \times 10^4 \text{ J}) \times \frac{1 \text{ photon}}{1.63 \times 10^{-24} \text{ J}} = \mathbf{3.3 \times 10^{28} \text{ photons}}$$

7.106 The Balmer series corresponds to transitions to the $n = 2$ level.

For He^+ :

$$\Delta E = R_{\text{He}^+} \left(\frac{1}{n_i^2} - \frac{1}{n_f^2} \right) \qquad \lambda = \frac{hc}{\Delta E} = \frac{(6.63 \times 10^{-34} \text{ J}\cdot\text{s})(3.00 \times 10^8 \text{ m/s})}{\Delta E}$$

For the transition, $n = 3 \rightarrow 2$

$$\Delta E = (8.72 \times 10^{-18} \text{ J}) \left(\frac{1}{3^2} - \frac{1}{2^2} \right) = -1.21 \times 10^{-18} \text{ J} \qquad \lambda = \frac{1.99 \times 10^{-25} \text{ J}\cdot\text{m}}{1.21 \times 10^{-18} \text{ J}} = 1.64 \times 10^{-7} \text{ m} = \mathbf{164 \text{ nm}}$$

For the transition, $n = 4 \rightarrow 2$, $\Delta E = -1.64 \times 10^{-18} \text{ J}$ $\lambda = \mathbf{121 \text{ nm}}$

For the transition, $n = 5 \rightarrow 2$, $\Delta E = -1.83 \times 10^{-18} \text{ J}$ $\lambda = \mathbf{109 \text{ nm}}$

For the transition, $n = 6 \rightarrow 2$, $\Delta E = -1.94 \times 10^{-18} \text{ J}$ $\lambda = \mathbf{103 \text{ nm}}$

For H, the calculations are identical to those above, except the Rydberg constant for H is $2.18 \times 10^{-18} \text{ J}$.

For the transition, $n = 3 \rightarrow 2$, $\Delta E = -3.03 \times 10^{-19} \text{ J}$ $\lambda = \mathbf{657 \text{ nm}}$

For the transition, $n = 4 \rightarrow 2$, $\Delta E = -4.09 \times 10^{-19} \text{ J}$ $\lambda = \mathbf{487 \text{ nm}}$

For the transition, $n = 5 \rightarrow 2$, $\Delta E = -4.58 \times 10^{-19} \text{ J}$ $\lambda = \mathbf{434 \text{ nm}}$

For the transition, $n = 6 \rightarrow 2$, $\Delta E = -4.84 \times 10^{-19} \text{ J}$ $\lambda = \mathbf{411 \text{ nm}}$

All the Balmer transitions for He^+ are in the ultraviolet region; whereas, the transitions for H are all in the visible region. Note the negative sign for energy indicating that a photon has been emitted.

7.107 (a) $\Delta H^\circ = \Delta H_f^\circ(\text{O}) + \Delta H_f^\circ(\text{O}_2) - \Delta H_f^\circ(\text{O}_3) = 249.4 \text{ kJ/mol} + (0) - 142.2 \text{ kJ/mol} = \mathbf{107.2 \text{ kJ/mol}}$

(b) The energy in part (a) is for *one mole* of photons. To apply $E = hv$ we must divide by Avogadro's number. The energy of one photon is:

$$E = \frac{107.2 \text{ kJ}}{1 \text{ mol}} \times \frac{1 \text{ mol}}{6.022 \times 10^{23} \text{ photons}} \times \frac{1000 \text{ J}}{1 \text{ kJ}} = 1.780 \times 10^{-19} \text{ J/photon}$$

The wavelength of this photon can be found using the relationship $E = hc/\lambda$.

$$\lambda = \frac{hc}{E} = \frac{(6.63 \times 10^{-34} \text{ J}\cdot\text{s})(3.00 \times 10^8 \text{ m/s})}{1.780 \times 10^{-19} \text{ J}} \times \frac{1 \text{ nm}}{1 \times 10^{-9} \text{ m}} = \mathbf{1.12 \times 10^3 \text{ nm}}$$

- 7.108** First, we need to calculate the energy of one 600 nm photon. Then, we can determine how many photons are needed to provide 4.0×10^{-17} J of energy.

The energy of one 600 nm photon is:

$$E = \frac{hc}{\lambda} = \frac{(6.63 \times 10^{-34} \text{ J}\cdot\text{s})(3.00 \times 10^8 \text{ m/s})}{600 \times 10^{-9} \text{ m}} = 3.32 \times 10^{-19} \text{ J/photon}$$

The number of photons needed to produce 4.0×10^{-17} J of energy is:

$$(4.0 \times 10^{-17} \text{ J}) \times \frac{1 \text{ photon}}{3.32 \times 10^{-19} \text{ J}} = \mathbf{1.2 \times 10^2 \text{ photons}}$$

- 7.109** Since the energy corresponding to a photon of wavelength λ_1 equals the energy of photon of wavelength λ_2 plus the energy of photon of wavelength λ_3 , then the equation must relate the wavelength to energy.

$$\text{energy of photon 1} = (\text{energy of photon 2} + \text{energy of photon 3})$$

Since $E = \frac{hc}{\lambda}$, then:

$$\frac{hc}{\lambda_1} = \frac{hc}{\lambda_2} + \frac{hc}{\lambda_3}$$

Dividing by hc :

$$\frac{1}{\lambda_1} = \frac{1}{\lambda_2} + \frac{1}{\lambda_3}$$

- 7.110** A “blue” photon (shorter wavelength) is higher energy than a “yellow” photon. For the same amount of energy delivered to the metal surface, there must be fewer “blue” photons than “yellow” photons. Thus, the yellow light would eject more electrons since there are more “yellow” photons. Since the “blue” photons are of higher energy, blue light will eject electrons with greater kinetic energy.

- 7.111** Energy of a photon at 360 nm:

$$E = h\nu = \frac{hc}{\lambda} = \frac{(6.63 \times 10^{-34} \text{ J}\cdot\text{s})(3.00 \times 10^8 \text{ m/s})}{360 \times 10^{-9} \text{ m}} = 5.53 \times 10^{-19} \text{ J}$$

Area of exposed body in cm^2 :

$$0.45 \text{ m}^2 \times \left(\frac{1 \text{ cm}}{1 \times 10^{-2} \text{ m}} \right)^2 = 4.5 \times 10^3 \text{ cm}^2$$

The number of photons absorbed by the body in 2 hours is:

$$0.5 \times \frac{2.0 \times 10^{16} \text{ photons}}{\text{cm}^2 \cdot \text{s}} \times (4.5 \times 10^3 \text{ cm}^2) \times \frac{7200 \text{ s}}{2 \text{ hr}} = 3.2 \times 10^{23} \text{ photons/2 hr}$$

The factor of 0.5 is used above because only 50% of the radiation is absorbed.

3.2×10^{23} photons with a wavelength of 360 nm correspond to an energy of:

$$(3.2 \times 10^{23} \text{ photons}) \times \frac{5.53 \times 10^{-19} \text{ J}}{1 \text{ photon}} = \mathbf{1.8 \times 10^5 \text{ J}}$$

7.112 The wavelength of a He atom can be calculated using the de Broglie equation. First, we need to calculate the root-mean-square speed using Equation (5.16) from the text.

$$u_{\text{rms}} = \sqrt{\frac{3 \left(8.314 \frac{\text{J}}{\text{K} \cdot \text{mol}} \right) (273 + 20) \text{K}}{4.003 \times 10^{-3} \text{ kg/mol}}} = 1.35 \times 10^3 \text{ m/s}$$

To calculate the wavelength, we also need the mass of a He atom in kg.

$$\frac{4.003 \times 10^{-3} \text{ kg He}}{1 \text{ mol He}} \times \frac{1 \text{ mol He}}{6.022 \times 10^{23} \text{ He atoms}} = 6.647 \times 10^{-27} \text{ kg/atom}$$

Finally, the wavelength of a He atom is:

$$\lambda = \frac{h}{mu} = \frac{(6.63 \times 10^{-34} \text{ J} \cdot \text{s})}{(6.647 \times 10^{-27} \text{ kg})(1.35 \times 10^3 \text{ m/s})} = \mathbf{7.39 \times 10^{-11} \text{ m} = 7.39 \times 10^{-2} \text{ nm}}$$

7.113 As an estimate, we can equate the energy for ionization ($\text{Fe}^{13+} \rightarrow \text{Fe}^{14+}$) to the average kinetic energy ($\frac{3}{2}RT$) of the ions.

$$\frac{3.5 \times 10^4 \text{ kJ}}{1 \text{ mol}} \times \frac{1000 \text{ J}}{1 \text{ kJ}} = 3.5 \times 10^7 \text{ J}$$

$$IE = \frac{3}{2}RT$$

$$3.5 \times 10^7 \text{ J/mol} = \frac{3}{2}(8.314 \text{ J/mol} \cdot \text{K})T$$

$$\mathbf{T = 2.8 \times 10^6 \text{ K}}$$

The actual temperature can be, and most probably is, higher than this.

7.114 The energy given in the problem is the energy of 1 mole of gamma rays. We need to convert this to the energy of one gamma ray, then we can calculate the wavelength and frequency of this gamma ray.

$$\frac{1.29 \times 10^{11} \text{ J}}{1 \text{ mol}} \times \frac{1 \text{ mol}}{6.022 \times 10^{23} \text{ gamma rays}} = 2.14 \times 10^{-13} \text{ J/gamma ray}$$

Now, we can calculate the wavelength and frequency from this energy.

$$E = \frac{hc}{\lambda}$$

$$\lambda = \frac{hc}{E} = \frac{(6.63 \times 10^{-34} \text{ J} \cdot \text{s})(3.00 \times 10^8 \text{ m/s})}{2.14 \times 10^{-13} \text{ J}} = \mathbf{9.29 \times 10^{-13} \text{ m} = 0.929 \text{ pm}}$$

and

$$E = h\nu$$

$$\nu = \frac{E}{h} = \frac{2.14 \times 10^{-13} \text{ J}}{6.63 \times 10^{-34} \text{ J}\cdot\text{s}} = 3.23 \times 10^{20} \text{ s}^{-1}$$

- 7.115** (a) **False.** $n = 2$ is the first excited state.
 (b) **False.** In the $n = 4$ state, the electron is (on average) further from the nucleus and hence easier to remove.
 (c) **True.**
 (d) **False.** The $n = 4$ to $n = 1$ transition is a higher energy transition, which corresponds to a *shorter* wavelength.
 (e) **True.**

7.116 Based on the *selection rule*, which states that $\Delta l = \pm 1$, only (b) and (d) are allowed transitions.

- 7.117** (a) Line A corresponds to the longest wavelength or lowest energy transition, which is the $3 \rightarrow 2$ transition. Therefore, line B corresponds to the $4 \rightarrow 2$ transition, and line C corresponds to the $5 \rightarrow 2$ transition.
 (b) We can derive an equation for the energy change (ΔE) for an electronic transition.

$$E_f = -R_H Z^2 \left(\frac{1}{n_f^2} \right) \quad \text{and} \quad E_i = -R_H Z^2 \left(\frac{1}{n_i^2} \right)$$

$$\Delta E = E_f - E_i = -R_H Z^2 \left(\frac{1}{n_f^2} \right) - \left(-R_H Z^2 \left(\frac{1}{n_i^2} \right) \right)$$

$$\Delta E = R_H Z^2 \left(\frac{1}{n_i^2} - \frac{1}{n_f^2} \right)$$

Line C corresponds to the $5 \rightarrow 2$ transition. The energy change associated with this transition can be calculated from the wavelength (27.1 nm).

$$E = \frac{hc}{\lambda} = \frac{(6.63 \times 10^{-34} \text{ J}\cdot\text{s})(3.00 \times 10^8 \text{ m/s})}{(27.1 \times 10^{-9} \text{ m})} = 7.34 \times 10^{-18} \text{ J}$$

For the $5 \rightarrow 2$ transition, we now know ΔE , n_i , n_f , and R_H ($R_H = 2.18 \times 10^{-18} \text{ J}$). Since this transition corresponds to an emission process, energy is released and ΔE is negative. ($\Delta E = -7.34 \times 10^{-18} \text{ J}$). We can now substitute these values into the equation above to solve for Z .

$$\Delta E = R_H Z^2 \left(\frac{1}{n_i^2} - \frac{1}{n_f^2} \right)$$

$$-7.34 \times 10^{-18} \text{ J} = (2.18 \times 10^{-18} \text{ J}) Z^2 \left(\frac{1}{5^2} - \frac{1}{2^2} \right)$$

$$-7.34 \times 10^{-18} \text{ J} = (-4.58 \times 10^{-19} \text{ J}) Z^2$$

$$Z^2 = 16.0$$

$$Z = 4$$

Z must be an integer because it represents the atomic number of the parent atom.

Now, knowing the value of Z , we can substitute in n_i and n_f for the $3 \rightarrow 2$ (Line A) and the $4 \rightarrow 2$ (Line B) transitions to solve for ΔE . We can then calculate the wavelength from the energy.

For Line A ($3 \rightarrow 2$)

$$\begin{aligned}\Delta E &= R_{\text{H}}Z^2 \left(\frac{1}{n_i^2} - \frac{1}{n_f^2} \right) = (2.18 \times 10^{-18} \text{ J})(4)^2 \left(\frac{1}{3^2} - \frac{1}{2^2} \right) \\ \Delta E &= -4.84 \times 10^{-18} \text{ J} \\ \lambda &= \frac{hc}{E} = \frac{(6.63 \times 10^{-34} \text{ J}\cdot\text{s})(3.00 \times 10^8 \text{ m/s})}{(4.84 \times 10^{-18} \text{ J})} = \mathbf{4.11 \times 10^{-8} \text{ m} = 41.1 \text{ nm}}\end{aligned}$$

For Line B ($4 \rightarrow 2$)

$$\begin{aligned}\Delta E &= R_{\text{H}}Z^2 \left(\frac{1}{n_i^2} - \frac{1}{n_f^2} \right) = (2.18 \times 10^{-18} \text{ J})(4)^2 \left(\frac{1}{4^2} - \frac{1}{2^2} \right) \\ \Delta E &= -6.54 \times 10^{-18} \text{ J} \\ \lambda &= \frac{hc}{E} = \frac{(6.63 \times 10^{-34} \text{ J}\cdot\text{s})(3.00 \times 10^8 \text{ m/s})}{(6.54 \times 10^{-18} \text{ J})} = \mathbf{3.04 \times 10^{-8} \text{ m} = 30.4 \text{ nm}}\end{aligned}$$

(c) The value of the final energy state is $n_f = \infty$. Use the equation derived in part (b) to solve for ΔE .

$$\begin{aligned}\Delta E &= R_{\text{H}}Z^2 \left(\frac{1}{n_i^2} - \frac{1}{n_f^2} \right) = (2.18 \times 10^{-18} \text{ J})(4)^2 \left(\frac{1}{4^2} - \frac{1}{\infty^2} \right) \\ \Delta E &= \mathbf{2.18 \times 10^{-18} \text{ J}}\end{aligned}$$

(d) As we move to higher energy levels in an atom or ion, the energy levels get closer together. See Figure 7.9 of the text, which represents the energy levels for the hydrogen atom. Transitions from higher energy levels to the $n = 2$ level will be very close in energy and hence will have similar wavelengths. The lines are so close together that they overlap, forming a continuum. The continuum shows that the electron has been removed from the ion, and we no longer have quantized energy levels associated with the electron. In other words, the energy of the electron can now vary continuously.

7.118 To calculate the energy to remove an electron from the $n = 1$ state and the $n = 5$ state in the Li^{2+} ion, we use the equation derived in Problem 7.117 (b).

$$\Delta E = R_{\text{H}}Z^2 \left(\frac{1}{n_i^2} - \frac{1}{n_f^2} \right)$$

For $n_i = 1$, $n_f = \infty$, and $Z = 3$, we have:

$$\Delta E = (2.18 \times 10^{-18} \text{ J})(3)^2 \left(\frac{1}{1^2} - \frac{1}{\infty^2} \right) = \mathbf{1.96 \times 10^{-17} \text{ J}}$$

For $n_i = 5$, $n_f = \infty$, and $Z = 3$, we have:

$$\Delta E = (2.18 \times 10^{-18} \text{ J})(3)^2 \left(\frac{1}{5^2} - \frac{1}{\infty^2} \right) = \mathbf{7.85 \times 10^{-19} \text{ J}}$$

To calculate the wavelength of the emitted photon in the electronic transition from $n = 5$ to $n = 1$, we first calculate ΔE and then calculate the wavelength.

$$\Delta E = R_H Z^2 \left(\frac{1}{n_i^2} - \frac{1}{n_f^2} \right) = (2.18 \times 10^{-18} \text{ J})(3)^2 \left(\frac{1}{5^2} - \frac{1}{1^2} \right) = -1.88 \times 10^{-17} \text{ J}$$

We ignore the minus sign for ΔE in calculating λ .

$$\lambda = \frac{hc}{\Delta E} = \frac{(6.63 \times 10^{-34} \text{ J}\cdot\text{s})(3.00 \times 10^8 \text{ m/s})}{1.88 \times 10^{-17} \text{ J}}$$

$$\lambda = 1.06 \times 10^{-8} \text{ m} = 10.6 \text{ nm}$$

- 7.119 (a)** First, we need to calculate the moving mass of the proton, and then we can calculate its wavelength using the de Broglie equation.

$$m_{\text{moving}} = \frac{m_{\text{rest}}}{\sqrt{1 - \left(\frac{u}{c}\right)^2}} = \frac{1.673 \times 10^{-27} \text{ kg}}{\sqrt{1 - \left(\frac{(0.50)(3.00 \times 10^8 \text{ m/s})}{3.00 \times 10^8 \text{ m/s}}\right)^2}} = 1.93 \times 10^{-27} \text{ kg}$$

$$\lambda = \frac{h}{mu} = \frac{6.63 \times 10^{-34} \text{ J}\cdot\text{s}}{(1.93 \times 10^{-27} \text{ kg})[(0.50)(3.00 \times 10^8 \text{ m/s})]}$$

$$\lambda = 2.29 \times 10^{-15} \text{ m} = 2.29 \times 10^{-6} \text{ nm}$$

(b)
$$m_{\text{moving}} = \frac{m_{\text{rest}}}{\sqrt{1 - \left(\frac{u}{c}\right)^2}} = \frac{6.0 \times 10^{-2} \text{ kg}}{\sqrt{1 - \left(\frac{63 \text{ m/s}}{3.00 \times 10^8 \text{ m/s}}\right)^2}} \approx 6.0 \times 10^{-2} \text{ kg}$$

The equation is only important for speeds close to that of light. Note that photons have a rest mass of zero; otherwise, their moving mass would be infinite!

- 7.120** We calculate W (the energy needed to remove an electron from the metal) at a wavelength of 351 nm. Once W is known, we can then calculate the velocity of an ejected electron using light with a wavelength of 313 nm.

First, we convert wavelength to frequency.

$$\nu = \frac{c}{\lambda} = \frac{3.00 \times 10^8 \text{ m/s}}{351 \times 10^{-9} \text{ m}} = 8.55 \times 10^{14} \text{ s}^{-1}$$

$$h\nu = W + \frac{1}{2}m_e u^2$$

$$(6.63 \times 10^{-34} \text{ J}\cdot\text{s})(8.55 \times 10^{14} \text{ s}^{-1}) = W + \frac{1}{2}(9.1094 \times 10^{-31} \text{ kg})(0 \text{ m/s})^2$$

$$W = 5.67 \times 10^{-19} \text{ J}$$

Next, we convert a wavelength of 313 nm to frequency, and then calculate the velocity of the ejected electron.

$$\nu = \frac{c}{\lambda} = \frac{3.00 \times 10^8 \text{ m/s}}{313 \times 10^{-9} \text{ m}} = 9.58 \times 10^{14} \text{ s}^{-1}$$

$$h\nu = W + \frac{1}{2}m_e u^2$$

$$(6.63 \times 10^{-34} \text{ J}\cdot\text{s})(9.58 \times 10^{14} \text{ s}^{-1}) = (5.67 \times 10^{-19} \text{ J}) + \frac{1}{2}(9.1094 \times 10^{-31} \text{ kg})u^2$$

$$6.82 \times 10^{-20} = (4.5547 \times 10^{-31})u^2$$

$$u = 3.87 \times 10^5 \text{ m/s}$$

- 7.121 (a)** We note that the maximum solar radiation centers around 500 nm. Thus, over billions of years, organisms have adjusted their development to capture energy at or near this wavelength. The two most notable cases are photosynthesis and vision.
- (b)** Astronomers record blackbody radiation curves from stars and compare them with those obtained from objects at different temperatures in the laboratory. Because the shape of the curve and the wavelength corresponding to the maximum depend on the temperature of an object, astronomers can reliably determine the temperature at the surface of a star from the closest matching curve and wavelength.