Sample Exercise 6.1 Concepts of Wavelength and Frequency

Two electromagnetic waves are represented in the margin. (**a**) Which wave has the higher frequency? (**b**) If one wave represents visible light and the other represents infrared radiation, which wave is which?

Solution

(a) The lower wave has a longer wavelength (greater distance between peaks). The longer the wavelength, the lower the frequency ($\nu = c/\lambda$). Thus, the lower wave has the lower frequency, and the upper wave has the higher frequency.

(b) The electromagnetic spectrum (Figure 6.4) indicates that infrared radiation has a longer wavelength than visible light. Thus, the lower wave would be the infrared radiation.

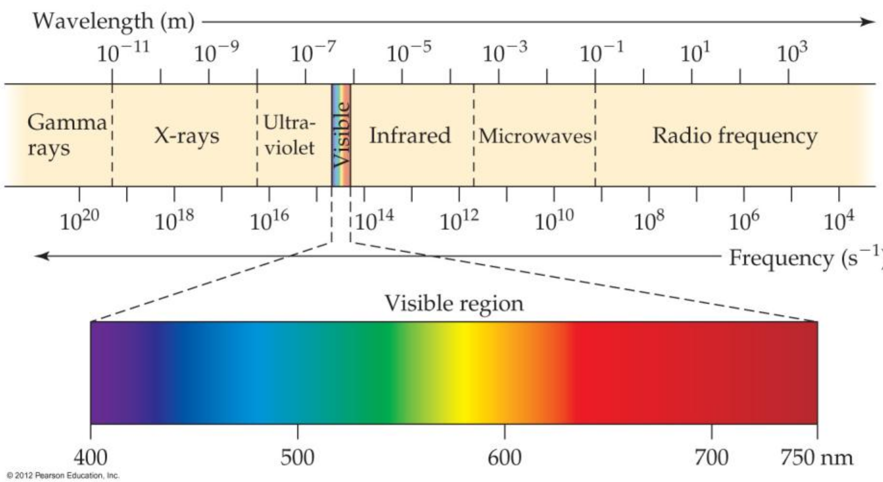

FIGURE 6.4 The electromagnetic spectrum. Wavelengths in the spectrum range from very short gamma rays to very long radio waves.

Practice Exercise

If one of the waves in the margin represents blue light and the other red light, which is which?

Answer: The expanded visible-light portion of Figure 6.4 tells you that red light has a longer wavelength than blue light. The lower wave has the longer wavelength (lower frequency) and would be the red light..

Sample Exercise 6.2 Calculating Frequency from Wavelength

The yellow light given off by a sodium vapor lamp used for public lighting has a wavelength of 589 nm. What is the frequency of this radiation?

Solution

Analyze We are given the wavelength, λ, of the radiation and asked to calculate its frequency, ν.

Plan The relationship between the wavelength and the frequency is given by Equation 6.1.We can solve for ν and use the values of λ and *c* to obtain a numerical answer. (The speed of light, *c*, is a fundamental constant whose value is 3.00×10^8 m/s.)

Solve Solving Equation 6.1 for frequency gives $\nu = c/\lambda$. When we insert the values for *c* and λ, we note that the units of length in these two quantities are different. We can convert the wavelength from nanometers to meters, so the units cancel:

Check The high frequency is reasonable because of the short wavelength. The units are proper because frequency has units of "per second," or s^{-1}.

Practice Exercise

(**a**) A laser used in eye surgery to fuse detached retinas produces radiation with a wavelength of 640.0 nm. Calculate the frequency of this radiation. (**b**) An FM radio station broadcasts electromagnetic radiation at a frequency of 103.4 MHz (megahertz; 1 MHz = $10^6 s^{-1}$). Calculate the wavelength of this radiation. The speed of light is 2.998×10^8 m/s to four significant digits.

Answers: (**a**) $4.688 \times 10^{14} s^{-1}$, (**b**) 2.899 m

Sample Exercise 6.3 Energy of a Photon

Calculate the energy of one photon of yellow light that has a wavelength of 589 nm.

Solution

Analyze Our task is to calculate the energy, E, of a photon, given $\lambda = 589$ nm.

Plan We can use Equation 6.1 to convert the wavelength to frequency: $\nu = c/\lambda$

We can then use Equation 6.3 to calculate energy: $E = h\nu$

Solve The frequency, ν, is calculated from the given wavelength, as shown in Sample Exercise 6.2: $\nu = c/\lambda = 5.09 \times 10^{14} s^{-1}$

The value of Planck's constant, h, is given both in the text and in the table of physical constants on the inside back cover of the text, and so we can easily calculate E:

$$E = (6.626 \times 10^{-34}\ \text{J-s})(5.09 \times 10^{14}\ s^{-1}) = 3.37 \times 10^{-19}\ \text{J}$$

Comment If one photon of radiant energy supplies 3.37×10^{-19} J, then one mole of these photons will supply

$$(6.02 \times 10^{23}\ \text{photons/mol})(3.37 \times 10^{-19}\text{J/photon}) = 2.03 \times 10^{5}\text{J/mol}$$

Practice Exercise

(**a**) A laser emits light that has a frequency of $4.69 \times 10^{14} s^{-1}$. What is the energy of one photon of this radiation? (**b**) If the laser emits a pulse containing 5.0×10^{17} photons of this radiation, what is the total energy of that pulse? (**c**) If the laser emits 1.3×10^{-2}J of energy during a pulse, how many photons are emitted?

Answers: (**a**) 3.11×10^{-19} J, (**b**) 0.16 J, (**c**) 4.2×10^{16} photons

Sample Exercise 6.4 Electronic Transitions in the Hydrogen Atom

Using Figure 6.12, predict which of these electronic transitions produces the spectral line having the longest wavelength: $n = 2$ to $n = 1$, $n = 3$ to $n = 2$, or $n = 4$ to $n = 3$.

Solution

The wavelength increases as frequency decreases ($\lambda = c/\nu$). Hence, the longest wavelength will be associated with the lowest frequency. According to Planck's equation, $E = h\nu$, the lowest frequency is associated with the lowest energy. In Figure 6.12 the energy levels (horizontal lines) that are closest together represents the smallest energy change. Thus, the $n = 4$ to $n = 3$ transition produces the longest wavelength (lowest frequency) line.

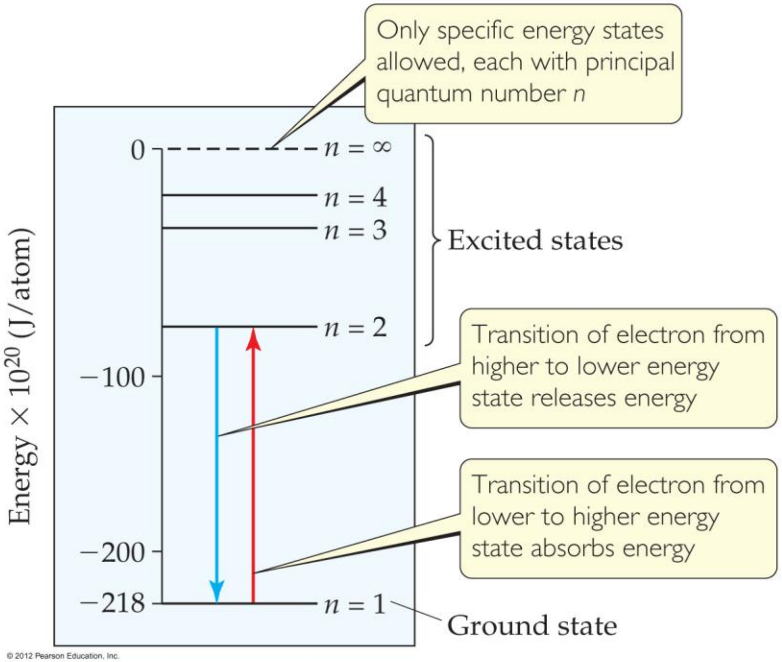

FIGURE 6.12 Energy states in the hydrogen atom. Only states for $n = 1$ through $n = 4$ and $n = \infty$ are shown. Energy is released or absorbed when an electron moves from one energy state to another.

Practice Exercise

Indicate whether each of the following electronic transitions emits energy or requires the absorption of energy: (**a**) $n = 3$ to $n = 1$; (**b**) $n = 2$ to $n = 4$.

Answers: (**a**) emits energy, (**b**) requires absorption of energy

Sample Exercise 6.5 Matter Waves

What is the wavelength of an electron moving with a speed of 5.97×10^6 m/s? The mass of the electron is 9.11×10^{-31} kg.

Solution

Analyze We are given the mass, *m*, and velocity, *v*, of the electron, and we must calculate its de Broglie wavelength, λ.
Plan The wavelength of a moving particle is given by Equation 6.8, so λ is calculated by inserting the known quantities *h*, *m*, and *v*. In doing so, however, we must pay attention to units.
Solve Using the value of Planck's constant, $h = 6.626 \times 10^{-34}$ J-s
We have the following:

Comment By comparing this value with the wavelengths of electromagnetic radiation shown in Figure 6.4, we see that the wavelength of this electron is about the same as that of X-rays.

Practice Exercise

Calculate the velocity of a neutron whose de Broglie wavelength is 500 pm. The mass of a neutron is given in the table inside the back cover of the text.

Answer: 7.92×10^2 m/s

Sample Exercise 6.6 Subshells of the Hydrogen Atom

(a) Without referring to Table 6.2, predict the number of subshells in the fourth shell, that is, for $n = 4$. (**b**) Give the label for each of these subshells. (**c**) How many orbitals are in each of these subshells?

Analyze and Plan We are given the value of the principal quantum number, n. We need to determine the allowed values of l and m_l for this given value of n and then count the number of orbitals in each subshell.

Solution

There are four subshells in the fourth shell, corresponding to the four possible values of l (0, 1, 2, and 3).
These subshells are labeled 4*s*, 4*p*, 4*d*, and 4*f*. The number given in the designation of a subshell is the principal quantum number, n; the letter designates the value of the angular momentum quantum number, l : for $l = 0$, *s*; for $l = 1$,*p*; for $l = 2$,*d*; for $l = 3$,*f*.

There is one 4*s* orbital (when $l = 0$, there is only one possible value of m_l: 0). There are three 4*p* orbitals (when $l = 1$, there are three possible values of m_l: 1, 0, -1). There are five 4*d* orbitals (when $1 = 2$, there are five allowed values of m_l: 2, 1, 0, -1, -2). There are seven 4*f* orbitals (when $l = 3$, there are seven permitted values of m_l: 3, 2, 1, 0, -1, -2, -3).

Practice Exercise

(**a**) What is the designation for the subshell with $n = 5$ and $l = 1$? (**b**) How many orbitals are in this subshell?
(**c**) Indicate the values of m_l for each of these orbitals.

Answers: (**a**) 5*p*; (**b**) 3; (**c**) 1, 0, -1

Sample Exercise 6.7 Orbital Diagrams and Electron Configurations

Draw the orbital diagram for the electron configuration of oxygen, atomic number 8. How many unpaired electrons does an oxygen atom possess?

Solution

Analyze and Plan Because oxygen has an atomic number of 8, each oxygen atom has 8 electrons. Figure 6.24 shows the ordering of orbitals. The electrons (represented as arrows) are placed in the orbitals (represented as boxes) beginning with the lowest-energy orbital, the 1*s*. Each orbital can hold a maximum of two electrons (the Pauli exclusion principle). Because the 2*p* orbitals are degenerate, we place one electron in each of these orbitals (spin-up) before pairing any electrons (Hund's rule).

Solve Two electrons each go into the 1*s* and 2*s* orbitals with their spins paired. This leaves four electrons for the three degenerate 2*p* orbitals. Following Hund's rule, we put one electron into each 2*p* orbital until all three orbitals have one electron each. The fourth electron is then paired up with one of the three electrons already in a 2*p* orbital, so that the orbital diagram is:

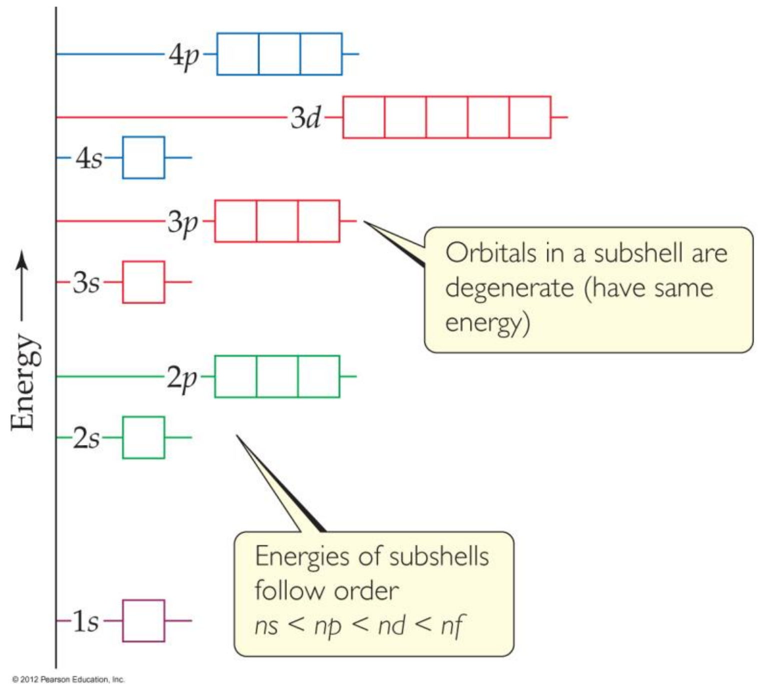

FIGURE 6.24 General energy ordering of orbitals for a many-electron atom.

Sample Exercise 6.7 Orbital Diagrams and Electron Configurations

Continued

The corresponding electron configuration is written $1s^22s^22p^4$. The atom has two unpaired electrons.

Practice Exercise

(**a**) Write the electron configuration for phosphorus, element 15. (**b**) How many unpaired electrons does a phosphorus atom possess?

Answers: (**a**) $1s^22s^22p^63s^23p^3$, (**b**) three

Sample Exercise 6.8 Electron Configurations for a Group

What is the characteristic valence electron configuration of the group 7A elements, the halogens?

Solution

Analyze and Plan We first locate the halogens in the periodic table, write the electron configurations for the first two elements, and then determine the general similarity between the configurations.

Solve The first member of the halogen group is fluorine (F, element 9). Moving backward from F, we find that the noble-gas core is [He]. Moving from He to the element of next higher atomic number brings us to Li, element 3. Because Li is in the second period of the *s* block, we add electrons to the 2*s* subshell. Moving across this block gives $2s^2$. Continuing to move to the right, we enter the *p* block. Counting the squares to F gives $2p^5$. Thus, the condensed electron configuration for fluorine is F: [He]$2s^22p^5$

The electron configuration for chlorine, the second halogen, is Cl: [Ne]$3s^23p^5$

From these two examples, we see that the characteristic valence electron configuration of a halogen is ns^2np^5, where *n* ranges from 2 in the case of fluorine to 6 in the case of astatine.

Practice Exercise

Which family of elements is characterized by an ns^2np^2 electron configuration in the outermost occupied shell?

Answer: group 4A

Sample Exercise 6.9 Electron Configurations from the Periodic Table

(**a**) Based on its position in the periodic table, write the condensed electron configuration for bismuth, element 83.
(**b**) How many unpaired electrons does a bismuth atom have?

Solution

(**a**) Our first step is to write the noble-gas core. We do this by locating bismuth, element 83, in the periodic table. We then move backward to the nearest noble gas, which is X*e*, element 54. Thus, the noble-gas core is [X*e*].

Next, we trace the path in order of increasing atomic numbers from X*e* to Bi. Moving from X*e* to Cs, element 55, we find ourselves in period 6 of the *s* block. Knowing the block and the period identifies the subshell in which we begin placing outer electrons, 6*s*. As we move through the s block, we add two electrons: $6s^2$.

As we move beyond the s block, from element 56 to element 57, the curved arrow below the periodic table reminds us that we are entering the *f* block. The first row of the *f* block corresponds to the 4*f* subshell. As we move across this block, we add 14 electrons: $4f^{14}$.

With element 71, we move into the third row of the *d* block. Because the first row of the *d* block is 3*d*, the second row is 4*d* and the third row is 5*d*.Thus, as we move through the ten elements of the *d* block, from element 71 to element 80, we fill the 5*d* subshell with ten electrons: $5d^{10}$.

Moving from element 80 to element 81 puts us into the *p* block in the 6*p* subshell. (Remember that the principal quantum number in the *p* block is the same as in the *s* block.)

Sample Exercise 6.9 Electron Configurations from the Periodic Table

Continued

Moving across to Bi requires 3 electrons: $6p^3$. The path we have taken is

Putting the parts together, we obtain the condensed electron configuration: $[Xe]6s^24f^{14}5d^{10}6p^3$. This configuration can also be written with the subshells arranged in order of increasing principal quantum number: $[Xe]4f^{14}5d^{10}6s^26p^3$.

Finally, we check our result to see if the number of electrons equals the atomic number of Bi, 83: Because Xe has 54 electrons (its atomic number), we have 54 + 2 + 14 + 10 + 3 = 83. (If we had 14 electrons too few, we would realize that we have missed the *f* block.)

(**b**) We see from the condensed electron configuration that the only partially occupied subshell is 6*p*. The orbital diagram representation for this subshell is

In accordance with Hund's rule, the three 6*p* electrons occupy the three 6*p* orbitals singly, with their spins parallel. Thus, there are three unpaired electrons in the bismuth atom.

Sample Exercise 6.9 Electron Configurations from the Periodic Table

Continued

Practice Exercise

Use the periodic table to write the condensed electron configuration for (**a**) Co (element 27), (**b**) T*e* (element 52).

Answers: (**a**) [Ar]$4s^2 3d^7$ or [Ar]$3d^7 4s^2$, (**b**) [Kr}$5s^2 4d^{10} 5p^4$ or [Kr]$4d^{10} 5s^2 5p^4$