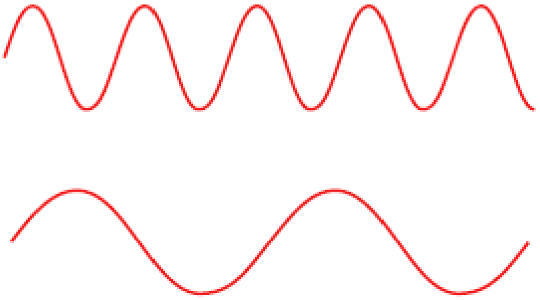
Sample Exercise 6.1 Concepts of Wavelength and Frequency

Two electromagnetic waves are represented below.



(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.

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, *v*.
Plan: The relationship between the wavelength (which is given) and the frequency (which is the unknown) is given by Equation 6.1. We can solve this equation for *v* and then use the values of 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 $v = c/\lambda$. 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:

$$\nu = \frac{c}{\lambda} = \left(\frac{3.00 \times 10^8 \text{ m/s}}{589 \text{ nm}}\right)\left(\frac{1 \text{ nm}}{10^{-9} \text{ m}}\right) = 5.09 \times 10^{14} \text{ s}^{-1}$$

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; MHz = 10^6 s^{-1}). Calculate the wavelength of this radiation. The speed of light is 2.998 × 108 m/s to four significant digits.
Answers: **(a)** 4.688×10^{14} s^{-1}, **(b)** 2.901 m

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Sample Exercise 6.3 Energy of a Photon

Calculate the energy of one photon of yellow light with 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, *v*, is calculated from the given wavelength, as shown in Sample Exercise 6.2:

$$\nu = c/\lambda = 5.09 \times 10^{14}\ \mathrm{s^{-1}}$$

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

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

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

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

This is the magnitude of enthalpies of reactions (Section 5.4), so radiation can break chemical bonds, producing what are called *photochemical reactions*.

Sample Exercise 6.3 Energy of a Photon

Practice Exercise

(a) A laser emits light with a frequency of 4.69×10^{14} s^{-1}. What is the energy of one photon of the radiation from this laser? **(b)** If the laser emits a pulse of energy 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 during the pulse?

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 6.14, predict which of the following 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 ($\lambda = c/v$). Hence the longest wavelength will be associated with the lowest frequency. According to Planck's equation, $E = hv$, the lowest frequency is associated with the lowest energy. In Figure 6.14 the shortest vertical line represents the smallest energy change. Thus, the $n = 4$ to $n = 3$ transition produces the longest wavelength (lowest frequency) line.

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 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,

$$h = 6.626 \times 10^{-34}\ \text{J-s}$$

and recalling that

$$1\ \text{J} = 1\ \text{kg-m}^2/\text{s}^2$$

we have the following:

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

$$= \frac{(6.626 \times 10^{-34}\ \text{J-s})}{(9.11 \times 10^{-31}\ \text{kg})(5.97 \times 10^6\ \text{m/s})}\left(\frac{1\ \text{kg-m}^2/\text{s}^2}{1\ \text{J}}\right)$$

$$= 1.22 \times 10^{-10}\ \text{m} = 0.122\ \text{nm} = 1.22\ \text{Å}$$

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.

Sample Exercise 6.5 Matter Waves

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, and –1). There are five 4*d* orbitals (when l = 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

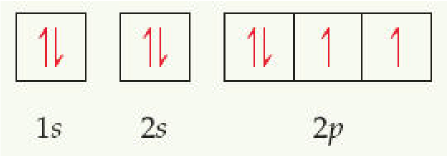
(a) What is the designation for the subshell with $n = 5$ and $l = 1$? **(b)** How many orbitals are in this subshell? **(c)** Indicate the values of m_l for each of these orbitals.
Answers: **(a)** 5*p*; **(b)** 3; **(c)** 1, 0, –1

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 6.25 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 representation is



The corresponding electron configuration is written $1s^2 2s^2 2p^4$. 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)** $1s^2 2s^2 2p^6 3s^2 3p^3$, **(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 them.
Solve: The first member of the halogen group is fluorine, atomic number 9. The condensed electron configuration for fluorine is

$$\text{F: [He]}2s^2 2p^5$$

Similarly, that for chlorine, the second halogen, is

$$\text{Cl: [Ne]}3s^2 3p^5$$

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) Write the electron configuration for bismuth, element number 83. **(b)** Write the condensed electron configuration for this element. **(c)** How many unpaired electrons does each atom of bismuth possess?

Solution

(a) We write the electron configuration by moving across the periodic table one row at a time and writing the occupancies of the orbital corresponding to each row (refer to Figure 6.29).

First row	$1s^2$
Second row	$2s^22p^6$
Third row	$3s^23p^6$
Fourth row	$4s^23d^{10}4p^6$
Fifth row	$5s^24d^{10}5p^6$
Sixth row	$6s^24f^{14}5d^{10}6p^3$
Total:	$1s^22s^22p^63s^23p^63d^{10}4s^24p^64d^{10}4f^{14}5s^25p^65d^{10}6s^26p^3$

Note that 3 is the lowest possible value that *n* may have for a *d* orbital and that 4 is the lowest possible value of *n* for an *f* orbital.
The total of the superscripted numbers should equal the atomic number of bismuth, 83. The electrons may be listed, as shown above in the "Total" row, in the order of increasing principal quantum number. However, it is equally correct to list the orbitals in the order in which they are read from Figure 6.30: $1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^24d^{10}5p^66s^24f^{14}5d^{10}6p^3$.
(b) We write the condensed electron configuration by locating bismuth on the periodic table and then moving *backward* to the nearest noble gas, which is Xe, element 54. Thus, the noble-gas core is [Xe]. The outer electrons are then read from the periodic table as before. Moving from Xe to Cs, element 55, we find ourselves in the sixth row. Moving across this row to Bi gives us the outer electrons. Thus, the abbreviated electron configuration is $[Xe]6s^24f^{14}5d^{10}6p^3$ or $[Xe]4f^{14}5d^{10}6s^26p^3$.

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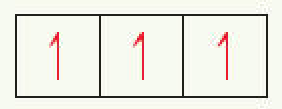
Sample Exercise 6.9 Electron Configurations from the Periodic Table

Solution (continued)

(c) We can see from the abbreviated electron configuration that the only partially occupied subshell is the 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 each atom of bismuth.

Practice Exercise

Use the periodic table to write the condensed electron configurations for **(a)** Co (atomic number 27) **(b)** Te (atomic number 52).

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

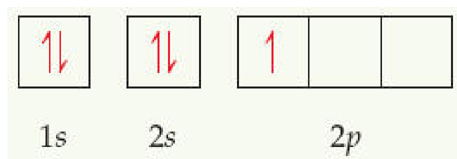
Sample Integrative Exercise Putting Concepts Together

Boron, atomic number 5, occurs naturally as two isotopes, ^{10}B and ^{11}B, with natural abundances of 19.9% and 80.1%, respectively. **(a)** In what ways do the two isotopes differ from each other? Does the electronic configuration of ^{10}B differ from that of ^{11}B? **(b)** Draw the orbital diagram for an atom of ^{11}B. Which electrons are the valence electrons? **(c)** Indicate three major ways in which the 1*s* electrons in boron differ from its 2*s* electrons. **(d)** Elemental boron reacts with fluorine to form BF_3, a gas. Write a balanced chemical equation for the reaction of solid boron with fluorine gas. **(e)** ΔH°_f for $BF_3(g)$ is –1135.6 kJ mol^{-1}. Calculate the standard enthalpy change in the reaction of boron with fluorine. **(f)** When BCl_3, also a gas at room temperature, comes into contact with water, the two react to form hydrochloric acid and boric acid, H_3BO_3, a very weak acid in water. Write a balanced net ionic equation for this reaction.

Solution

(a) The two isotopes of boron differ in the number of neutrons in the nucleus. (Sections 2.3 and 2.4) Each of the isotopes contains five protons, but ^{10}B contains five neutrons, whereas ^{11}B contains six neutrons. The two isotopes of boron have identical electron configurations, $1s^2 2s^2 2p^1$, because each has five electrons.

(b) The complete orbital diagram is



The valence electrons are the ones in the outermost occupied shell, the $2s^2$ and $2p^1$ electrons. The $1s^2$ electrons constitute the core electrons, which we represent as [He] when we write the condensed electron configuration, $[He]2s^2 2p^1$.

(c) The 1*s* and 2*s* orbitals are both spherical, but they differ in three important respects: First, the 1*s* orbital is lower in energy than the 2*s* orbital. Second, the average distance of the 2*s* electrons from the nucleus is greater than that of the 1*s* electrons, so the 1*s* orbital is smaller than the 2*s*. Third, the 2*s* orbital has one node, whereas the 1*s* orbital has no nodes (Figure 6.19).

Sample Integrative Exercise Putting Concepts Together

Solution (continued)

(d) The balanced chemical equation is

$$2\ B(s) + 3\ F_2(g) \rightarrow 2\ BF_3(g)$$

(e) $\Delta H° = 2(-1135.6) - [0 + 0] = -2271.2$ kJ. The reaction is strongly exothermic.
(f) $BCl_3(g) + 3\ H_2O(l) \rightarrow H_3BO_3(aq) + 3\ H^+(aq) + 3\ Cl^-(aq)$. Note that because H_3BO_3 is a very weak acid, its chemical formula is written in molecular form, as discussed in Section 4.3.