Solutions for Chapter 4 End-of-Chapter Problems

Problem 4.1.
(a) Mendeleev could not have based his periodic table on the structure of atoms, because the structure of atoms was not understood until considerably after the time of Mendeleev. The basis for understanding the present periodic arrangement of atoms, the number of protons in the nucleus, was not experimental evidence available to Mendeleev in 1869.
(b) Mendeleev’s periodic table was based on empirical correlations among the physical and chemical properties of the elements and their compounds. The experimental properties of the 60 elements known at that time showed repeated patterns based on their relative atomic masses.
(c) To align known elements with similar chemical properties in families, Mendeleev had to leave gaps in the first periodic table. Mendeleev assumed that the gaps would be filled by elements that had not yet been discovered and he made predictions about the properties of these yet-to-be-discovered elements.

Problem 4.2.
The data for this problem are in this table.

<table>
<thead>
<tr>
<th>Element</th>
<th>Radius (pm)</th>
<th>Element</th>
<th>Radius (pm)</th>
<th>Element</th>
<th>Radius (pm)</th>
</tr>
</thead>
<tbody>
<tr>
<td>H</td>
<td>37</td>
<td>N</td>
<td>75</td>
<td>Al</td>
<td>118</td>
</tr>
<tr>
<td>He</td>
<td>32</td>
<td>O</td>
<td>73</td>
<td>Si</td>
<td>111</td>
</tr>
<tr>
<td>Li</td>
<td>134</td>
<td>F</td>
<td>71</td>
<td>P</td>
<td>105</td>
</tr>
<tr>
<td>Be</td>
<td>90</td>
<td>Ne</td>
<td>69</td>
<td>S</td>
<td>102</td>
</tr>
<tr>
<td>B</td>
<td>82</td>
<td>Na</td>
<td>154</td>
<td>Cl</td>
<td>99</td>
</tr>
<tr>
<td>C</td>
<td>77</td>
<td>Mg</td>
<td>130</td>
<td>Ar</td>
<td>97</td>
</tr>
</tbody>
</table>

(a) We see that there is a large jump in the atomic radius going from He to Li and then eight elements later another large jump going from Ne to Na. In each case, the atomic radii of the following elements decreases as the atomic number increases until the next jump in size occurs. On the basis of the data here, we predict that there will be another jump in radius going from Ar to the next element, K, since these are another eight elements after the second jump.
(b) An obvious way to plot these data is a graph similar to Figure 4.1 or 4.2 in the text, but with the atomic number (or equal spacing between elements) as the independent variable, as in Figure 4.3. Here is such a plot on which the periodic pattern is quite obvious. (The points are connected only to help visualize the trends; this is not a continuous plot but a series of discrete points with no values between them.)
Another possible way to display the data graphically is to show the relative sizes of the atoms as circles of radii proportional to the atomic radii. Such a graphical array in a pattern that is like a conventional periodic table is shown here:

(c) It is almost always easier to visualize trends from a well-made graphical representation than trying to pick out the trends from a table of data. Exact numerical values are more easily found from a table.

(d) The atomic radius of K will be larger than that of Na, just as Na is larger than Li, since the size of the atoms appears to increase down a family of the periodic table. If we look at the differences between the sizes of the atoms in the second and third periods on our plots (or from the table), we see that third period radii are 30-40 pm larger than the second period radii. In these data, we see that the difference in radii between Ne and H is 37 pm and between Ar and Ne is 28 pm. We might conclude that the difference between the radii of K (fourth period) and Na (third period) will be somewhere in this range and would then predict that the radius of K would be about 35 pm larger than that for Na, or about 189 pm. A similar analysis for Be, Mg, and Ca leads to a prediction of approximately 165 pm for the radius of Ca.

(e) The data in the table come from the WebElements web site, www.webelements.com, where the values for the covalent atomic radii of K and Ca are 196 pm and 174 pm, respectively. Our predictions in part (d) are in the ballpark, although about 5% low. This relatively good agreement shows the usefulness of periodic trends to make estimates of elemental atomic properties. The data for most elements are on this plot (a screen capture from the web site):
Problem 4.3.
(a) The energy required to remove a second electron from helium should be considerably greater than that required to remove the first electron. Removing the second electron leaves behind a doubly positively charged ion, He$^{2+}$. The attraction the electron feels as it is removed is at least twice that for the removal of the first electron, which leaves behind only a singly charged ion, He$^+$. Also, when the first electron is close to the nucleus, it is repelled by the second, and this makes it easier to remove the first electron. The experimental data are:

First ionization energy:  He (g) → He$^+$ (g) + e$^-$ 2372 kJ·mol$^{-1}$
Second ionization energy:  He$^+$ (g) → He$^{2+}$ (g) + e$^-$ 7300 kJ·mol$^{-1}$

(b) The second ionization energy for beryllium should be greater than the first because the electron is being removed from the positively charged ion, leaving behind a doubly positively charged ion, as in part (a) for helium. The third ionization energy should be much greater, because this electron is coming from the electron core, not the outer valence electrons.

First ionization energy:  Be (g) → Be$^+$ (g) + e$^-$ 900 kJ·mol$^{-1}$
Second ionization energy:  Be$^+$ (g) → Be$^{2+}$ (g) + e$^-$ 1760 kJ·mol$^{-1}$
Third ionization energy:  Be$^{2+}$ (g) → Be$^{3+}$ (g) + e$^-$ 14,850 kJ·mol$^{-1}$

Problem 4.4.
This problem asks you to compare Figure 4.4 with this representation of first ionization energies for the representative (main group) elements (families labeled I through VIII).
Although we may be more familiar with the sort of graphical representation in Figure 4.4, this three-dimensional representation has some added advantages. One of the main advantages is the positioning of the families (groups) that makes it easier to see the relationships within families as well as within periods. This is particularly important in this representation that shows main group elements for six periods, not just the first three, as in Figure 4.4. Both representations show the change in ionization energy within each period and group. It is easier to read numeric values from the two-dimensional graph in Figure 4.4, but it may be easier to observe some apparent irregularities in the patterns with the three-dimensional graph. For example, why is the value for nitrogen so much higher than its neighbors within the second period and within group V? Both representations show the shell structure of atoms and this structure falls naturally out of the two-dimensional plot when the ionization energies are plotted as a function of atomic number. However, in the three-dimensional representation, the repetitive structure (the substantial drops in ionization energies after each noble gas) is imposed on the data in order to position the third dimension (the families) appropriately. Thus, the repeating (shell) structure is used in the construction, rather than appearing naturally.

**Problem 4.5.**

The data for this problem are in this table.

<table>
<thead>
<tr>
<th>Element</th>
<th>$IE_2$ (kJ·mol$^{-1}$)</th>
</tr>
</thead>
<tbody>
<tr>
<td>H</td>
<td>0</td>
</tr>
<tr>
<td>He</td>
<td>5248</td>
</tr>
<tr>
<td>Li</td>
<td>7300</td>
</tr>
<tr>
<td>Be</td>
<td>1757</td>
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<tr>
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<td>2430</td>
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<td>2350</td>
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<td>N</td>
<td>2860</td>
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<td>O</td>
<td>3390</td>
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<td>F</td>
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<td>Ne</td>
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</tr>
<tr>
<td>K</td>
<td>3052</td>
</tr>
<tr>
<td>Ca</td>
<td>1145</td>
</tr>
</tbody>
</table>

(a) There is no value for the second ionization energy of the hydrogen atom, because the atom has only one electron; there is not a second one to lose.
(b) A plot of the second ionization energy data in the table (together with the data for the first ionization energy from Figure 4.4 and Table 4.2) as a function of atomic number gives:

![Graph showing second ionization energy data as a function of atomic number.](image)

Comparison of the plots of the first and second ionization energies shows that the shape of the two curves is essentially identical, except that the second ionization energy plot is moved over one atomic number higher and the energies on this plot are a good deal higher. It looks like the ionization energies of the mononegative ions from He\(^{+}\) through Ca\(^{+}\), follow exactly the same pattern as the first ionization energies for H through K. Each ion in this series of ions has the same number of electrons as the corresponding atom in the other series. That is, He\(^{+}\) has one electron and H has one electron and so on to Ca\(^{+}\) and K, each of which has 19 electrons. The shell structure applies to ions as well as atoms and depends on the number of electrons surrounding the nucleus. The second ionization energies for an elemental atom are all higher than the first, because the electron is being removed from a positive ion and is held more tightly. (Compare with the solution to Problem 4.3.)

**Problem 4.6.**

(a) Three adjacent pairs of atomic masses that are “out of order” and one pair that are almost the same are shown on this screen capture of a plot of molar mass as a function of atomic number from the interactive periodic table in the *Web Companion*. The modern relative atomic masses are shown for each element. (Other pairs at higher atomic number are also “out of order,” but in these cases, one or both of the elements does not exist naturally, so only the isotopes that have been prepared synthetically are available. We do not know what a “natural” distribution of isotopes would be, nor, of course, did Mendeleev know about any of these.)
(b) Since Mendeleev ordered the elements by relative atomic mass (or molar mass), he would have problems with the three pairs that are “out of order,” because their chemical properties would not “fit” in the family in which their mass would have put them. This was not a problem for the Ar-K pair, since Ar (and the other noble gases) had not been discovered when Mendeleev created his first table. Note that Mendeleev assigned both of the Co-Ni pair the same relative atomic mass; these were probably the best data available in 1869. For the Te-I pair, Mendeleev gives the correct relative atomic masses (he uses “J,” the German symbol for iodine, since he published in a German journal), but reverses the order given by the atomic mass, in order to put the elements in the correct chemical families (iodine with the other halogens). He indicates his concern about this by placing a question mark after the value of Te, presumably to indicate that this value might be suspect.

Problem 4.7.
Sodium metal is a solid at room temperature. The metal melts at 97.8 °C and boils at 882.9 °C. It is not necessary to heat the sample to its boiling point to do the experiment represented in Figure 4.6(b); a vapor pressure less than one atmosphere provides sufficient gaseous atoms to observe the absorption spectrum shown in the figure, so the sample only needs to be heated to a few hundred degrees.

Problem 4.8.
Atoms strongly heated at the sun’s surface, like the atoms in an electrical discharge tube in Figure 4.6(a), are given a great deal of energy that they lose by emitting light. Atoms further from the sun are not heated strongly enough to make them emit, but they can absorb light energy from the sun, just as the sample of atoms in Figure 4.6(b) does.

Problem 4.9.
(a), (b), (c) See Figure 4.8 for a representation of a wave with the wavelength, amplitude, and node identified.
(d) A cycle is one complete traversal of the wave plot from some point to the first identical point further along the plot. In Figure 4.8, for example, the adjacent maximum amplitudes that denote the wavelength also denote one full cycle of the wave. The entire wave is composed of repeated cycles.

Problem 4.10.
(a) Because, even at its closest approach to Earth, Mars is 56 million km from Earth, it takes time for a radio message, traveling at the speed of light, $3.00 \times 10^5$ m·s$^{-1}$, from a space probe on Mars to reach the Earth. The time required is:
time = (56 * 10^6 km) * \(\frac{1000 \text{ m}}{1 \text{ km}}\) * \(\frac{1 \text{ s}}{3.00 \times 10^8 \text{ m}}\) = 187 s (a little more than 3 minutes)

(b) Controllers on Earth have problems maneuvering a remote-controlled vehicle on the Martian surface because it takes a minimum of about three minutes to get a signal from the vehicle and another three minutes for a command to reach the vehicle. This is one of the reasons that the vehicles are designed to move so slowly. The more recent rovers, Spirit and Opportunity that landed in early 2004, are more sophisticated than the 1997 Sojourner Rover, Figure 1.39, and do some of their own computer-controlled navigation (with less immediate guidance from the Earth-based human controllers).

Problem 4.11.
For each of the frequencies, \(\nu\), calculate the wavelength, \(\lambda\), using the relationship \(c/\nu\), where \(c\) is the speed of light, \(3.00 \times 10^8 \text{ m} \cdot \text{s}^{-1}\), and Figure 4.13 to identify the type of radiation (region of the electromagnetic spectrum).

(a) 101 MHz:
\[
\nu = \frac{3.00 \times 10^8 \text{ m} \cdot \text{s}^{-1}}{101 \times 10^6 \text{ s}^{-1}} = 297 \text{ m} \quad \text{(AM radio)}
\]

(b) 25 \(\times 10^{15}\) Hz
\[
\nu = \frac{3.00 \times 10^8 \text{ m} \cdot \text{s}^{-1}}{25 \times 10^{15} \text{ s}^{-1}} = 1.2 \times 10^{-8} \text{ m} = 12 \text{ nm} \quad \text{(ultraviolet)}
\]

(c) 200 GHz
\[
\nu = \frac{3.00 \times 10^8 \text{ m} \cdot \text{s}^{-1}}{200 \times 10^9 \text{ s}^{-1}} = 1.5 \times 10^{-3} \text{ m} = 1.5 \text{ mm} \quad \text{(microwave)}
\]

Problem 4.12.
For each of the wavelengths, \(\lambda\), calculate the frequency, \(\nu\), using the relationship \(c/\lambda\), where \(c\) is the speed of light, \(3.00 \times 10^8 \text{ m} \cdot \text{s}^{-1}\), and Figure 4.13 to identify the type of radiation (region of the electromagnetic spectrum).

(a) 400 nm
\[
\nu = \frac{3.00 \times 10^8 \text{ m} \cdot \text{s}^{-1}}{400 \times 10^{-9} \text{ m}} = 7.5 \times 10^{-14} \text{ s}^{-1} \quad \text{(visible)}
\]

(b) 25 km
\[
\nu = \frac{3.00 \times 10^8 \text{ m} \cdot \text{s}^{-1}}{25 \times 10^3 \text{ m}} = 1.2 \times 10^4 \text{ s}^{-1} \quad \text{(radio, very low frequency)}
\]

(c) 4.5 \(\times 10^{-4}\) m
\[
\nu = \frac{3.00 \times 10^8 \text{ m} \cdot \text{s}^{-1}}{4.5 \times 10^{-4} \text{ m}} = 6.7 \times 10^{11} \text{ s}^{-1} \quad \text{(radar, UHF-TV)}
\]

Problem 4.13.
(a) A wave with a wavelength of 0.34 m and a frequency of 0.75 s\(^{-1}\) has a speed, \(u\), of:
\[
u = \frac{u}{\lambda} = (0.75 \text{ s}^{-1})(0.34 \text{ m}) = 0.26 \text{ m} \cdot \text{s}^{-1}
\]

(b) This cannot be a wave of electromagnetic radiation, because electromagnetic waves travel at the speed of light, \(3.00 \times 10^8 \text{ m} \cdot \text{s}^{-1}\).

The wavelength of microwave radiation with a frequency of 1.145 \(\times 10^{10}\) s\(^{-1}\) is:
\[ c = \frac{2.998 \times 10^8 \text{ m s}^{-1}}{1.145 \times 10^{10} \text{ s}^{-1}} = 0.02598 \text{ m} = 25.98 \text{ mm} \]

**Problem 4.15.**
The frequency corresponding to an ultraviolet wavelength of 280 nm is:
\[ c = \frac{3.00 \times 10^8 \text{ m s}^{-1}}{280 \times 10^{-9} \text{ m}} = 1.07 \times 10^{15} \text{ s}^{-1} \]

**Problem 4.16.**
The length of a crossbar in an antenna that receives a signal with a frequency of $3 \times 10^3 \text{ MHz}$ should have a length that is approximately the same as the wavelength of this signal. The wavelength is:
\[ c = \frac{3.00 \times 10^8 \text{ m s}^{-1}}{3 \times 10^8 \text{ s}^{-1}} = 1 \text{ m} \]
The crossbars on TV antennas are 1 to 2 m long, so this answer is in the right ballpark for a VHF-TV antenna.

**Problem 4.17.**
(a) A screen capture from the *Web Companion*, Chapter 4, Section 4.2, page 4 (with the correct set of oscillating arrows representing the electromagnetic waves chosen) is shown here.

[NOTE: The oscillating arrows in the Companion are slightly misleading because the blue arrow is 180° out of phase with the representation of the wave. That is, when the blue electric field wave is moving up, as in this screen capture, the blue arrow in the box is pointing down and *vice versa*. In this picture, the screen capture has been manipulated to flip the blue arrow, so that it corresponds to the position of the motion of the blue wave in the electromagnetic wave representation.] The feedback box in the Companion tells you how the correct set of oscillating arrows is related to the representation of the electromagnetic wave. The electric and magnetic waves oscillate perpendicular to one another, as shown by the arrows. At any instant (as captured here) the amplitudes of the two waves are the same, as shown by the equal lengths of the arrows.
(b) The description of electromagnetic radiation given in this chapter of the text is essentially identical to that outlined in part (a). Thus, the description of the relationship between the oscillating arrows in the Companion and the text description, including Figure 4.12 (which is essentially identical to this one in the Companion), is the same as in part (a).

Problem 4.18.
Figure 4.14 shows that the wave theory (which associates wave amplitude with energy) predicts emission from glowing bodies that is more intense than is observed and does not drop off at shorter wavelengths.

Problem 4.19.
(a) The results of photoelectric effect experiments showed that the intensity of light did not affect the maximum energy of the electrons being ejected. The maximum energy is a function of the wavelength of the incident light. The number of electrons ejected is a function of the intensity of the incident light. These effects are consistent with a particle description of light in which each particle carries an amount of energy that is determined by the wavelength and the intensity of the light is associated with the number of particles. Each particle (with sufficient energy, that is, short enough wavelength) can eject an electron whose energy is determined by the energy of the particle. The more particles there are (higher intensity), the more electrons can be ejected.

(b) The wave theory predicted that, if a light beam shining on a metal surface were bright enough (had a large enough amplitude or intensity), any frequency (wavelength) of light could eject an electron from the metal surface. The experimental results however do not support this prediction, because electrons are ejected only by light of a frequency higher than some minimum that is dependent on the metal being investigated.

Problem 4.20.
When Einstein analyzed the photoelectric experimental data quantitatively, he found that the maximum energy of the ejected electrons could be expressed in terms of the frequency of the incident light, \( \omega \), and a threshold frequency, \( \omega_0 \), as \( E_{\text{max}} = (\text{constant})(\omega - \omega_0) \). The numeric value of the constant in this relationship was identical to the numeric value of the constant, \( h \), that Planck had had to use to relate the frequency and energy of light to explain the emission from glowing hot objects. This finding reinforced the Planck hypothesis that light is quantized and comes in discrete energy packets, \( E = h \). Einstein used the term photon to describe the energy packets of light. Photons of higher-frequency radiation have higher energies, and photons of lower-frequency radiation have lower energy. If the photons of the radiation in a photoelectric experiment have less energy than the energy required for an electron to overcome the attractive forces that hold it within the metal, the electron cannot escape from the metal. If the incident photon energy is higher than the threshold energy for ejection, \( h \omega > h \omega_0 \), the maximum energy of the ejected electron is the difference between the incident and threshold energy, \( h \omega - h \omega_0 \).

Problem 4.21.
Two examples of quantization in everyday life are:

- monetary systems such as that used in the United States where all values are based on the “quantum” of one cent (a penny). [NOTE: For some financial transactions and computations, the system is further subdivided into mils, 1 mil = 0.1 cent, but there is no object that represents a mil, so for actual monetary transactions, the cent is the smallest quantum unit.]
- the notes of musical instruments constrained to only certain values that depend upon the length of strings (pianos and guitars) or a resonant pipe (flutes and trombones).

What others did you think of?
Problem 4.22.
The energy of a photon with a frequency of $1.255 \times 10^6 \text{ s}^{-1}$ is:

$$E = h \nu = (6.626 \times 10^{-34} \text{ J} \cdot \text{s})(1.255 \times 10^6 \text{ s}^{-1}) = 8.316 \times 10^{-28} \text{ J}$$

Problem 4.23.
(a) The frequency of the radiation associated with one photon of microwave radiation that has an energy of $1.64 \times 10^{-24} \text{ J}$ is:

$$= \frac{E}{h} = \frac{1.64 \times 10^{-24} \text{ J}}{6.626 \times 10^{-34} \text{ J} \cdot \text{s}} = 2.48 \times 10^9 \text{ s}^{-1}$$

(b) To compare the frequency of the microwave photon in part (a) with the frequency of green light with a wavelength of 515 nm, we have to calculate the frequency of the green light:

$$= \frac{3.00 \times 10^8 \text{ m} \cdot \text{s}^{-1}}{515 \times 10^{-9} \text{ m}} = 5.83 \times 10^{14} \text{ s}^{-1}$$

This frequency is approximately $10^5$ or 100,000 times higher than the frequency of the microwave radiation. Note: This frequency could also be found from the energy of green light calculated in Check This 4.33(a) from the relationship $E = h \nu$.

(c) Frequency is directly proportional to energy, so the photons of green light have about $10^5$ or 100,000 times more energy than photons of microwaves.

Problem 4.24.
(a) The wavelength (in meters) of the radiation used in a 500-MHz nuclear magnetic resonance (NMR) spectrometer is:

$$= \frac{c}{\nu} = \frac{3.00 \times 10^8 \text{ m} \cdot \text{s}^{-1}}{500 \times 10^6 \text{ s}^{-1}} = 0.600 \text{ m} = 600 \text{ mm}$$

(b) This NMR instrument operates in approximately the cellular phone wavelength region of the electromagnetic spectrum.

(c) The energy (in kJ·mol$^{-1}$) corresponding to a frequency of 500 MHz is:

$$E = h N_A = (6.626 \times 10^{-34} \text{ J} \cdot \text{s})(500 \times 10^6 \text{ s}^{-1})(6.022 \times 10^{23} \text{ mol}^{-1}) \frac{1 \text{ kJ}}{1000 \text{ J}}$$

$$= 2.00 \times 10^{-4} \text{ kJ}$$

Problem 4.25.
The energy of one photon and one mole of photons with
(a) a frequency of 101 MHz is:

$$E \text{ (per photon)} = h \nu = (6.626 \times 10^{-34} \text{ J} \cdot \text{s})(101 \times 10^6 \text{ s}^{-1}) = 6.69 \times 10^{-26} \text{ J}$$

$$E \text{ (per mole)} = h N_A = (6.69 \times 10^{-26} \text{ J})(6.022 \times 10^{23} \text{ mol}^{-1}) = 4.03 \times 10^{-2} \text{ J}$$

(b) a wavelength of 400 nm is:

$$E \text{ (per photon)} = \frac{hc}{\lambda} = (6.626 \times 10^{-34} \text{ J} \cdot \text{s})(3.00 \times 10^8 \text{ m} \cdot \text{s}^{-1})/(400 \times 10^{-9} \text{ m})$$

$$= 4.97 \times 10^{-19} \text{ J}$$

$$E \text{ (per mole)} = \frac{hc}{\lambda} N_A = (4.97 \times 10^{-19} \text{ J})(6.022 \times 10^{23} \text{ mol}^{-1}) = 2.99 \times 10^5 \text{ J} = 299 \text{ kJ}$$
The energy of a quantum of light with a wavelength of 589 nm (yellow light) is:

\[ E = \frac{hc}{\lambda} = \frac{(6.626 \times 10^{-34} \text{ J} \cdot \text{s})(3.00 \times 10^8 \text{ m/s})}{(589 \times 10^{-9} \text{ m})} = 3.37 \times 10^{-19} \text{ J} \]

The student’s calculation that gave 4.2 \times 10^{-19} \text{ J} was not correct. Although different amounts of energy can be emitted by the light, depending upon the number of photons emitted, these energies must all be multiples of 3.37 \times 10^{-19} \text{ J}. The light source cannot lose energy except in multiples of 3.37 \times 10^{-19} \text{ J} and the student's answer is not a multiple of 3.37 \times 10^{-19} \text{ J}.

Problem 4.27.
The lenses in continuous “photo-gray” or “photo-sun” eyeglasses contain silver chloride crystals evenly distributed in a glass or highly crystalline plastic matrix. As light hits the crystal, it provides the energy to transfer an electron from chloride ion to silver ion, forming tiny specks of silver metal that cause the lens to darken. The chlorine atom is not free to join other chlorine atoms and escape as gaseous chlorine because it is still locked within the crystalline matrix. This reaction is reversible, so in the absence of light, silver metal and chlorine atoms rejoin to form silver chloride crystals, which do not block light. The lens regains its former transparent appearance.

\[ \text{Ag}^+ \text{Cl}^- \xrightarrow{\text{light energy}} \text{Ag} \text{ (s)} + \text{Cl}^-(aq) \]

silver metal  chlorine atom

Problem 4.28.
In Investigate This 4.29, we found that red light does not cause the formation of Ag metal in a crystal of silver chloride, whereas blue and green light do cause this transformation. A red safelight is used in a photographic darkroom to prevent exposure of the film to light that will cause further reaction in the film being developed.

Problem 4.29.
In a Science article, scientists reported variations in the concentrations of thorium and potassium on the moon’s surface. They used gamma ray spectrometry that recorded the gamma ray emission from potassium at approximately 1.4 MeV and from thorium at approximately 2.6 MeV. Calculate the frequencies of these two gamma ray emissions. Calculate the wavelengths of these two gamma ray emissions. Note: 1 MeV = 1 \times 10^6 \text{ eV} (electron volt) and 1 eV = 1.602 \times 10^{-19} \text{ J}.

Answer to 4.29:
To find the frequencies and wavelengths of gamma ray emissions from potassium (emission energy of approximately 1.4 MeV) and thorium (emission energy of approximately 2.6 MeV), we need to convert the energies to joules and use the Planck relationship to get the frequency and the relationship among frequency, wavelength, and speed of light to calculate the wavelength.

For potassium, the energy conversion and calculations are:

\[ E = 1.4 \text{ MeV} = (1.4 \text{ MeV}) \frac{1 \text{ MeV}}{1 \times 10^6 \text{ eV}} \frac{1.602 \times 10^{-19} \text{ J}}{1 \text{ eV}} = 2.2 \times 10^{-13} \text{ J} \]

\[ \frac{E}{h} = \frac{2.2 \times 10^{-13} \text{ J}}{6.626 \times 10^{-34} \text{ J} \cdot \text{s}} = 3.4 \times 10^{20} \text{ s}^{-1} \]
c/ = \frac{3.00 \times 10^8 \text{ m s}^{-1}}{3.4 \times 10^{20} \text{ s}^{-1}} = 8.8 \times 10^{-13} \text{ m}

For thorium, the energy conversion and calculations are:

\[
E = 2.6 \text{ MeV} = (2.6 \text{ MeV}) \frac{1 \times 10^6 \text{ eV}}{1 \text{ MeV}} \frac{1.602 \times 10^{-19} \text{ J}}{1 \text{ eV}} = 4.2 \times 10^{-13} \text{ J}
\]

\[
\frac{E}{\hbar} = \frac{4.2 \times 10^{-13} \text{ J}}{6.626 \times 10^{-34} \text{ J s}} = 6.3 \times 10^{20} \text{ s}^{-1}
\]

\[
c/ = \frac{3.00 \times 10^8 \text{ m s}^{-1}}{6.3 \times 10^{20} \text{ s}^{-1}} = 4.8 \times 10^{-13} \text{ m}
\]

**Problem 4.30.**

(a) The spectrum of radiation emitted by a substance is called an emission spectrum. When the emitted light is from atoms that have been given energy (excited) by heating or by an electric discharge, only certain wavelengths are present in the spectrum. The emission spectrum consists of several sharp lines. In an excited atom some of the electrons are in higher energy levels of the atom and when they lose energy to go to lower energy levels, the difference in energy between the two levels is emitted in the form of a photon of light with a frequency/wavelength corresponding to this energy. Since only certain energy levels are allowed in the atom, the atomic emissions occur only at discrete energies (frequencies/wavelengths) corresponding to the differences in energy between the energy levels in the atom.

(b) When continuous electromagnetic radiation, such as that from a light bulb, passes through a substance, some wavelengths may be absorbed. The spectrum of the radiation that passes through is called an absorption spectrum. The absorption spectrum from gaseous atoms consists of what looks like a continuous spectrum with several black lines where light of these wavelengths has been absorbed by the atoms in the sample. The electrons in the atoms absorb energy and go from lower to higher energy levels. The wavelengths of the absorbed light corresponds to the photons with energies equal to the energy spacing between the lower and higher energy levels. Since only certain energy levels are allowed in the atom, the atomic absorptions occur only at certain energies (frequencies/wavelengths).

(c) For a given element, the atomic absorption and emission spectra are complementary. The wavelengths present in the emission spectrum are the same as those missing in the absorption spectrum. In the emission and absorption processes, light of exactly the same wavelength is emitted (when an electron changes from a higher to a lower energy level) or absorbed (when an electron changes from the lower to the higher energy level).

**Problem 4.31.**

The intensity of a spectral line in an atomic emission spectrum can be directly related to (iii) the number of electrons undergoing the transition that gives rise to the line. If more electrons undergo a transition, more photons will be emitted and a more intense line will be observed. Only two energy levels are involved in the transition that gives rise to an emission, so phrase (i) cannot be appropriate. The speed of transition between, phrase (ii), has no effect on the intensity of the emission. Indeed, the concept of the speed of the transition may be meaningless. If the electron took a finite amount of time to go from the upper to lower state, it would presumably have different amounts of energy, depending on how far along it was in the transition, and we might expect to observe emissions of different amounts of energy (more than one wavelength). As atoms are probed experimentally at shorter and shorter times, now approaching attoseconds, 1attosecond = 10^{-18} seconds, we may discover new phenomena associated with atomic energy.
Problem 4.32.
When electrons are excited from a ground state to an excited state, (iii) energy is absorbed. The excited state is a state with higher energy. An electron needs to absorb energy in order to reach this state. Phrases (i) and (iv) both have to do with emission of energy, which involves transitions from an excited state to a lower energy state (sometimes the ground state), so these are not appropriate to complete the sentence. Energy in the form of heat might be responsible for the excitation of an electron from the ground to the excited state, but the heat energy would have to be absorbed, not released, as phrase (ii) suggests.

Problem 4.33.
Excited atoms of neon gas inside glass tubing produce the familiar red-orange glow of many “neon” lights. The electrons in the neon atoms are excited by an electrical discharge. The atoms emit light when the high-energy, excited electrons return to lower energy levels. When the emitted light is dispersed by a prism, only certain wavelength are observed. These wavelengths correspond to the differences in energy between the atomic energy levels. Since only certain energy levels are allowed in the atoms, only certain frequencies of light are emitted and the spectrum is not continuous but consists of several sharp lines.

Problem 4.34.
When you add so much energy to an atom that an electron completely leaves the atom, it becomes an ion (a cation), since the electron is completely removed from the atom.

Problem 4.35.
(a) In a flame emission test, the yellow flame from a solution containing Na\(^+\) ions corresponds to a wavelength of about 600 nm while the violet flame from a solution containing K\(^+\) ions corresponds to a wavelength of about 400 nm. (See the visible electromagnetic spectrum in Check This 4.9.)
(b) Figure 4.6(a) represents the emission of light from an electrically excited sample of sodium gas. In a flame emission test the heat of the flame excites the atoms/ions in the sample. The energy levels of the atoms/ions are the same, no matter how they are excited, so emissions from the two sources are essentially the same. [More energy can usually be given the excited atoms in an electrical discharge, so, often, more lines are observed in the emission from a discharge than from a flame, but all the lines in the flame emission are also present in the discharge emission.]
(c) The emission from the K\(^+\) ion solution has the higher energy. Since energy is inversely proportional to wavelength, a lower wavelength corresponds to a higher energy.

Problem 4.36.
The colors to which the six wavelengths in the emission spectrum of mercury in the visible range correspond are: blue-violet to blue for the 405, 408, and 436 nm lines; green for the 546 nm line; and yellow-green for the 577 and 579 nm lines.
(b) Since the energy of an emission is inversely proportional to the wavelength of the emission, the 405 nm wavelength, the shortest wavelength, corresponds to the largest difference in energy between the atomic states of mercury responsible for the emission.
(c) The 579 nm wavelength, the longest wavelength, corresponds to the smallest difference in energy between the atomic states of mercury responsible for the emission.

Problem 4.37.
The line spectra of hydrogen, helium, mercury, and neon support the idea that the energy of electrons in atoms is quantized, because these discrete emission (or absorption) wavelengths represent energy changes of the electrons in the atoms. If the energies in the atoms could take
any value, then the emissions would be at all wavelengths (that is, continuous). The fact that the observed emissions are at only certain wavelengths is consistent with an atomic model in which electrons can have only certain energies, that is, that their energies are quantized.

Problem 4.38.
The electronic energy transitions in this problem refer to the energy levels on this energy level diagram
(a) The transition $E_1 \rightarrow E_2$ requires absorption of a photon to increase the energy of the electron from the ground state to the first excited state.
(b) The transition $E_2 \rightarrow E_1$ results in emission of a photon as the electron energy decreases going from the first excited state to the ground state. The wavelength of the emitted photon is the same as the wavelength of the photon absorbed in part (a).
(c) The transition $E_4 \rightarrow E_3$ results in emission of a photon as the electron energy decreases going from the third to the second excited state.
(d) The transition $E_4 \rightarrow E_5$ requires absorption of a photon to increase the energy of the electron from the third to the fourth excited state.
(e) The transition $E_6 \rightarrow E_2$ results in emission of a photon as the electron energy decreases going from the fifth to the first excited state.
(f) The transition $E_6 \rightarrow E_1$ results in emission of a photon as the electron energy decreases going from the fifth excited state to the ground state.

To rank the relative wavelengths of emission and absorption (shortest to longest) we need to rank the energy changes from largest to smallest. We can do this by comparing the distances between levels on the energy level diagram. Going from the highest level, $E_6$, to the lowest level, $E_1$, as in (f) above, is obviously the largest energy change. The changes between energy levels $E_1$ and $E_2$, (a) and (b), are the same; this is the second largest energy change. The change between $E_6$ and $E_2$, (e), brackets the remaining two changes, so it must be the third largest energy change. The remaining two changes both involve energy level $E_4$. It is obvious from the diagram that the energy difference between $E_4$ and $E_3$ is larger than that between $E_4$ and $E_5$, so (c) represents the fourth largest energy change and (d) the fifth largest (or smallest) energy change. Thus, the relative wavelengths of emission and absorption (shortest to longest) are:

(f) < (a) = (b) < (e) < (c) < (d)

Problem 4.39.
In theory, any moving object gives rise to a characteristic matter wave and de Broglie’s equation can be applied to both macroscopic and submicroscopic (atomic/molecular level) objects. However, the wave properties become observable only for submicroscopic objects. The wavelengths of large moving objects, (like vehicles or balls), are so small that they can not be detected by any existing measuring device.

Problem 4.40.
We use the de Broglie wavelength equation to calculate the wavelength of a baseball of mass 0.5 kg traveling at 30 m·s⁻¹:

$$\lambda = \frac{\hbar}{mu} = \frac{6.626 \times 10^{-34} \text{ J s}}{0.5 \text{ kg \cdot 30 m \cdot s}^{-1}} = 4 \times 10^{-35} \text{ m}$$

The wavelength of the moving baseball is so short that it is undetectable by any measuring device, so any affect on its behavior is undetectable.
Problem 4.41.
We use the de Broglie wavelength equation to calculate the wavelength of an electron of mass $9.1 \times 10^{-31}$ kg traveling at $1.5 \times 10^6$ m·s⁻¹:

$$\lambda = \frac{h}{mu} = \frac{6.626 \times 10^{-34} \text{ J s}}{9.1 \times 10^{-31} \text{ kg} \times 1.5 \times 10^6 \text{ m s}^{-1}} = 4.9 \times 10^{-10} \text{ m} \approx 490 \text{ pm}$$

Atoms have sizes that are measured in picometers (pm) so this wavelength is significant compared to the size of an atom.

Problem 4.42.
In the macroscopic world, particles are massive enough that their wave characteristics (wavelengths), although presumably present, are undetectable and hence do not significantly affect their properties. At the atomic level, particles have very small masses and both wave and particle properties can be observed (depending upon the observations that are made) and are important in understanding the interactions among matter and energy.

Problem 4.43.
(a) The de Broglie wavelength of an alpha particle, mass $6.6 \times 10^{-27}$ kg, traveling at a velocity of $1.5 \times 10^7$ m·s⁻¹ is:

$$\lambda = \frac{h}{mu} = \frac{6.626 \times 10^{-34} \text{ J s}}{6.6 \times 10^{-27} \text{ kg} \times 1.5 \times 10^7 \text{ m s}^{-1}} = 6.7 \times 10^{-15} \text{ m} = 6.7 \text{ fm (femtometer)}$$

(b) This very short wavelength is in the electromagnetic spectral region associated with highly energetic gamma rays.

Problem 4.44.
The standing waves on a guitar string are responsible for the sound produced by the guitar. The waves responsible for the sound produced by other musical instruments are described here.

(a) The drumhead on drums is set in motion by the drumstick(s) and two-dimensional standing waves are set up on the head.

(b) As air is blown across the mouthpiece by the flutist, standing waves are set up inside the tube. The wavelength of these waves is controlled by the length of the tube that is closed, which the player controls by opening and closing the valves on the instrument. You can get the same effect by blowing across the top of a bottle with a narrow opening, such as a soft-drink bottle. The frequency of the sound you produce can be changed by changing the amount of liquid in the bottle (thus changing the size of the air column you are setting in motion.

(c) When a tuning fork is struck, its legs begin to vibrate. The standing waves are set up in the legs, very much like waves on a string, but a bit more complicated because the fork legs are thicker and only one end is fixed. The pitch of the fork depends upon its various dimensions (length, cross section of the legs, etc.)

Problem 4.45.
The contributions four scientists made to our understanding of atomic and electronic structure are briefly described here:

(a) Bohr developed the theory of quantized electron energy levels.

(b) Heisenberg developed the uncertainty principle, which states that, at very small (atomic) scale, we cannot accurately and simultaneously measure both the momentum and position of an electron.

(c) Einstein showed that the quantized electromagnetic radiation proposed by Planck also quantitatively explained the photoelectric effect.
(d) De Broglie developed the hypothesis that moving atomic-level objects behave like waves as well as particles and thus led to the dual wave/particle nature of matter as well as light.

Problem 4.46.
The Heisenberg uncertainty principle tells us that the more we know about the location of an electron, the less we know about its momentum and *vice versa*.

Problem 4.47.
Bohr proposed that the electrons in an atom orbit the nucleus in a well-defined path, so it would be possible to determine precisely and simultaneously both the position of an electron (from the radius of the orbit) and its momentum (from its kinetic energy). However, the Heisenberg uncertainty principle states that we cannot determine precisely and simultaneously both the position and momentum of an electron in an atom. The de Broglie wave model states that the electron in the atom has properties of a wave as well as of a particle. Since a wave extends in space it is not possible to determine exactly where the electron is located at a specific time, which is consistent with Heisenberg uncertainty principle, and overcomes this limitation of the Bohr model.

Problem 4.48.
We can calculate exactly the position and the momentum of a rolling ball, because its wave properties are unobservable (too massive and too slow) and it behaves as a particle. We cannot calculate the same for a moving electron because it is very light and exhibits observable wave properties that make it impossible to simultaneously and precisely determine its position and momentum.

Problem 4.49.
The size of a hydrogen atom is a function of the kinetic and potential energy of the electron wave and the nucleus. As the radius of the electron cloud decreases the kinetic energy of the electron increases, raising the energy of the system. At the same time, the potential energy of the system decreases as the distance between the nucleus and the electron decreases, which lowers the energy of the system. These opposite effects result in a minimum energy of the system (electron and proton) at an electron wave radius of about 60 pm, the radius of the hydrogen atom.

Problem 4.50.
To compare the kinetic energy of particle A (mass $1.56 \times 10^{-25}$ kg) and particle B (mass $4.25 \times 10^{-24}$ kg), each traveling at a velocity of $3.15 \times 10^5$ m s$^{-1}$, we need only compare their masses, because they are traveling at the same velocity. Particle B has a greater kinetic energy, $m u^2/2$, because it has more mass.

Problem 4.51.
The potential energy of an electron wave becomes lower as it gets closer to the nucleus, that is, as the wave becomes smaller.

Problem 4.52.
In Figure 4.27, the potential energy is negative because it is proportional to the product of the charge on the electron ($-1$) and the charge on the nucleus ($+1$ for hydrogen). These particles with unlike charge attract one another, so the potential energy of the system is lowered by their interaction. Kinetic energy, $m u^2/2$, is always positive, because mass is always positive and
the velocity of the electron is squared, so, even if the velocity is negative (motion in a negative direction in the coordinate system used to describe the system), the function $m \cdot u^2$ is positive.

Problem 4.53.
The red circle in the *Web Companion*, Chapter 4, Section 4.7, page 1, represents the spherical electron wave that surrounds and interacts with the nucleus. The energy of the spherical electron wave is inversely proportional to the square of the radius of the wave. As the electron gets closer to the nucleus (the wave gets smaller) the kinetic energy, represented by the red curve on the graph, rises. If kinetic energy were the only energy in the system, the electron and proton would not come together.

Problem 4.54.
Electron spin is the property of the electron that makes it behave as though it is a tiny magnet. Electron spin is responsible for electrons interacting with magnetic fields.

Problem 4.55.
In the experiment represented in Figure 4.28(a), the magnetic field does not affect the stream of helium atoms, because the spins of the two electrons in helium are opposite so their interactions with the field cancel each other out. The magnetic field does affect the stream of lithium atoms, Figure 4.28(b), because there are three electrons in the atom two of which are paired with one another; the third electron has no partner (is unpaired) and interacts with the magnetic field. Since half the atoms (on average) have the odd electron spin in one direction and the other half in the other direction, the interaction splits the stream of atoms into two streams, as shown.

Problem 4.56.
To tell whether atoms of an element have an odd number of electrons (or electrons that are not spin paired with one another), you might pass a stream of atoms of the element through a magnetic field like the one represented in Figure 4.28. Elements whose atoms have a single unpaired spin (an odd number of electrons) would split into two beams as the atoms with spin up and spin down electrons interact differently with the magnetic field. [Streams of atoms of elements whose atoms have more than one unpaired spin would be split into more than two streams by the magnetic field.]

Problem 4.57.
(a) A nitrogen atom has seven electrons. If three of the electrons have unpaired spins, that means the other four are spin-paired (two pairs of spin-paired electrons).
(b) The electrons in a ground state atom are described by the lowest energy orbitals possible, so we might conclude that the energies of spatial orbitals that describe the three unpaired electrons are all the same. If one of the orbitals had a lower energy, then we would expect two electrons with opposite spins to be described by that orbital and to give an atom with only one unpaired spin.

Problem 4.58.
The Pauli exclusion principle states that only two electrons of opposite spin can occupy exactly the same region of space (orbital) in an atom. For lithium, the two lowest-energy electrons with paired spins occupy the orbital (shell) next to nucleus. The third electron, (which has the same spin as one or the other of the first two electrons), cannot share the same space with them. It has to be described by an orbital at a larger distance from the nucleus, and therefore has a lower ionization energy than the first two electrons.

Problem 4.59.
The second ionization energy for an atom is always greater than the first ionization energy, because, after the first electron is removed, a positive ion, more protons than electrons, has been formed. To remove an electron from this positive ion requires more energy (because of more
electrostatic attraction) than to remove the first electron from a neutral atom. Another way to state this is to say that there is a greater effective nuclear attraction for the remaining electrons when one (or more) has been removed.

**Problem 4.60.**
(a) These ions and atom, Na\(^+\), Ne, F\(^-\), and O\(^2-\), all have ten electrons (two in the first shell and eight in the second). Atoms and ions of different elements that have the same number of electrons are called “isoelectronic.”
(b) Of the four species in part (a), Na\(^+\) has the highest ionization energy. In all four species, the ten electrons have the same distribution about the nucleus, that is, two in the first shell and eight in the second. Thus, Na\(^+\), with the most protons, has the highest nuclear charge and holds its ten electrons with greater attraction than the other three ions and atom.

**Problem 4.61.**
Ionization energies, atomic radii, and electronegativities all depend on the nuclear charge, which increases across a row of the periodic table. Thus, attraction for electrons are more strongly attracted to the nucleus across a row of the table.
(a) Ionization energy increases across a period, because the attraction for the electrons increases across a period and it becomes harder to remove electrons.
(b) Atomic radii decrease across a row of the periodic table, because the attraction for the electrons increases across a period and they are held closer to the nucleus.
(c) Electronegativity is a measure of the relative attraction of an atom for the shared electrons in a covalent bond with another atom. Electronegativity increases across a row because the nuclear charge increases and the atoms attract electrons more strongly, including those they share with other atoms.

**Problem 4.62.**
Ionization energies increase as you go from left to right in a row (period) on the periodic table. Across a period electrons are added to the same shell and are all about the same distance (on average) from the nucleus. Since the nuclear charge increases across a period, the electrons are all held more strongly and are closer to the nucleus (the shell gets smaller) across a period. Thus, the energy it takes to remove an electron (the ionization energy) increases across a period.

**Problem 4.63.**
Ionization energies decrease as you go from top to bottom in a group (family) on the periodic table. Down a family electrons in successive periods are added to shells that are successively larger and (on average) a greater distance (on average) from the nucleus. Although the nuclear charge increases across down a family, the “effective” nuclear charge remains about the same because the core charge on the elemental atoms in a family is the same. Since the electrons are held by about the same core charge but are a greater distance from the nucleus, the energy it takes to remove an electron (the ionization energy) decreases down a period.

**Problem 4.64.**
The number of valence electrons for the elements in the first three periods is given by the Roman numeral designation for its group (family).
(a) Oxygen is in group VI and has six valence electrons. Oxygen is atomic number 8, so has a total of eight electrons, two of which are core electrons in the first shell.
(b) Sodium is in group I and has one valence electron. Sodium is atomic number 11, so has a total of 11 electrons, 10 of which are core electrons in the first and second shells.
(c) Chlorine is in group VII and has seven valence electrons. Chlorine is atomic number 17, so has a total of 17 electrons, 10 of which are core electrons in the first and second shells.
Argon is in group VIII and has eight valence electrons. Argon is atomic number 18, so has a total of 18 electrons, 10 of which are core electrons in the first and second shells.

Problem 4.65.
This table of the empirical formulas of the simplest hydrides of the second period elements is the background for this problem.

<table>
<thead>
<tr>
<th>element</th>
<th>Li</th>
<th>Be</th>
<th>B</th>
<th>C</th>
<th>N</th>
<th>O</th>
<th>F</th>
<th>Ne</th>
</tr>
</thead>
<tbody>
<tr>
<td>hydride</td>
<td>LiH</td>
<td>BeH₂</td>
<td>BH₃</td>
<td>CH₄</td>
<td>NH₃</td>
<td>OH₂</td>
<td>FH_</td>
<td>none</td>
</tr>
</tbody>
</table>

(a) As we go across the second period, we see that the number of hydrogen atoms in the empirical formulas of the simplest hydrides increases from one to four and then decreases to one (or none, if we count the noble gas, neon, which has not yet been found to form compounds). The ability to form bonds with hydrogen seems to be a maximum at Group 4, the carbon family, and to decrease as we go to higher or lower atomic numbers in the period. The contrasts between LiH and FH (or HF, as it is usually written) help to understand a bit about this series of compounds. LiH is an ionic salt of the Li⁺ cation and H⁻ anion. HF is a weak acid that can react with water to form the H₃O⁺ cation and F⁻ anion. At the left end of the period, where electronegativity and ionization energy are low, the Li atom donates a good deal of its valence electron density to the H atom when they bond. At the right end of the period, where electronegativity and ionization energy are high, the F atom accepts a good deal of valence electron density from the H atom, making it possible for the proton to transfer to water to form H₃O⁺. Thus, the H atom is an electron acceptor or electron donor, depending on the atom to which it bonds, so the “hydrides” are not strictly comparable with one another. They are quite polar (in opposite directions) at the ends of the period. In the middle of the period, CH₄ is a hydrocarbon, which we have seen in previous chapters is nonpolar. Thus, the properties and formulas of the second period hydrides both follow trends that change toward the middle of the period. Note that the number of H atoms bonded to the atoms on the left of the period is equal to the number of valence electrons in the atom. The number of H atoms bonded to the atoms on the right of the period is equal to the number of electrons that are lacking to give the atom a completed shell of eight electrons. In the middle, carbon atoms match both these patterns.

(b) The H atom has one electron in its valence shell. This is the same as the Li atom. Both atoms bond with one H atom to form, respectively, H₂ and LiH. Thus, the H and Li atoms follow the same bonding pattern as we saw in the table and part (a). Since the electronegativities of the H atom and Li atom are more than one unit different, we do not expect the chemistry of the two hydrides to be similar. Indeed, H₂ is nonpolar and a gas at room temperature, while LiH is an ionic solid at room temperature. We can also argue that an H atom lacks one electron of having a filled electron shell, which is similar to the F atom. Both atoms bond with one H atom to form, respectively, H₂ and FH (HF). Thus, the H and F atoms follow the same bonding pattern as we saw in the table and part (a). Since the electronegativities of the H atom and F atom are more than one unit different, we do not expect the chemistry of the two hydrides to be similar. Indeed, H₂ is a nonpolar gas at room temperature, while HF is a polar gas that dissolves in water to form an acidic solution [part (a)]. Another way to look at the H atom is to consider that it is a half-filled electron shell, one electron in a shell that can accommodate two. This is quite like carbon with four electrons in a shell that can accommodate eight. Since the electronegativities of C and H are quite close, they form a nonpolar gaseous compound, CH₄, that is more similar to H₂ than either LiH or HF. These differences and similarities are the reason we have chosen to locate the element hydrogen in the group IV family of the periodic table, instead of the more conventional location in group I (or occasionally in group VII).

(c) The third period elements have electron shell structures similar to those in the second period (Figure 4.31), except that the electrons are not as tightly held (lower ionization energies) and the electronegativities are somewhat lower for atoms in the same families. Based on the similarity in electronic structure, we would predict that the empirical formulas of the simplest
hydrides of Al, Si, and P would be AlH₃, SiH₄, and PH₃. We would also predict that SiH₄ would be relatively nonpolar. These predictions are borne out by experimental observations.

**Problem 4.66.**

(a) The plot of first ionization energies from the interactive periodic table in the *Web Companion* is shown here. The saw tooth shape with peaks at noble gases and minima at the alkali metals are like the plots in Figures 4.4 and 4.30. The smaller features, for example, the two, three, three steps between the minima and the peaks are evident here as well. By extending the plot beyond element 20, we see other features, the 10-element plateaus evident in periods four (19-36) and five (37-54) and following a 14-element lower plateau in period six (55-86). At the end of the plateau, there is a drop in ionization energy at the first of the last six elements in the period (which also show the smaller stair step feature). The 14-element low plateau is also evident in period seven, but the rest of the period is incomplete.

(b) The plot of second-ionization energies is shown here. The saw tooth features and the smaller intermediate features noted above in part (a) are also evident here, except, as we found in the solution for Problem 4.5, moved one atomic number (element) higher. The shell structure of a monopositive elemental ion appears to be the same and behave the same as that of the neutral atom one atomic number lower (which has the same number of electrons). Since the nuclear charge of the ion is one higher, the actual ionization energies are higher. [Note the difference in the energy scales here and in part (a).]

(c) The plot of third ionization energies is shown here. Although the smaller features are becoming blurred, they are still present and we see that dipositive elemental ions have
ionization properties similar to the monopositive ions and neutral atoms with the same number of electrons. Since the nuclear charge on the dipositive ions are higher than on the corresponding monopositive ions or neutral atoms, the ionization energies are higher still, than those in parts (a) and (b).

Problem 4.67.
Was your classmate able to solve your periodicity problem based on the interactive periodic table in the Web Companion? Explain why or why not.

Problem 4.68.
[NOTE: This problem is probably misplaced and should go in Section 4.11, Wave Equations and Atomic Orbitals, or in the section General Problems. It is helpful to discuss its structure and problems with the s, p, d, f nomenclature for atomic orbitals.]
In this three-dimensional periodic table model of eight stacked, round wooden discs with the symbols for the elements on the discs, each of the second tier discs represents parts of two periods or rows in the electron shell model. For example, the top disc shown here has the last six elements from period two on the outer ring and the first two elements of period three on the inner ring. Similarly, the lower disc represents the rest of period three and beginning of period four. [The two small discs at the top of the stack contain the two elements of the first period (on the top disc) and the first two elements of the second period (on the lower disc).] The problem with this representation is that periods are not kept together and can be confusing, if you do not realize that different periods are represented on the same disc. What is required is to read from the outside toward the center to get the ordering of shells and subshells correct on each disc. In terms of the structure of the shell and subshell structure of the periodic table, the inner ring is the two-electron subshell in each shell (s-block elements), the second ring is the six-electron subshell (p-block elements), the third ring is the 10 electron subshell (d-block elements), and the fourth ring is the 14 electron subshell (f-block elements). (See Section 4.11 and the electronic configurations there for an explanation of this nomenclature.)
[NOTE: The "Periodic Roundtable" is an "invention" of Gary Katz (PhD biochemist), P.O. Box 156, Cabot, VT 05647. He sells it in a hardwood version for about $50. It has been reviewed in the Journal of Chemical Education, 2000, 77(2), 164, by Glen Rogers, from Allegheny College.]
Problem 4.69.
NOTE: This problem is probably misplaced and should go in Section 4.11, Wave Equations and Atomic Orbitals, or in the section General Problems. It is helpful to discuss its structure and problems with the s, p, d, f nomenclature for atomic orbitals.

(a) The central cylinder of the three-dimensional periodic table shown here (and with other views in the chapter opening illustration) corresponds to the elements in the two far left and six far right groups of the conventional periodic table. These elements are often called the representative elements. Often, the two columns of elements at the right are called the s-block elements and the six on the right, the p-block elements. (See Section 4.11 and the electronic configurations there for an explanation of this nomenclature.) The broader loop is attached to the central cylinder between the second and third columns of the representative elements, beginning at period four, that is, between the alkaline earth metals and the boron-family elements. This is where the interrupting block of elements (10 per period) is found in the conventional periodic table. This block of elements is often called the transition metals or the d-block transition metals. The thinner loop is attached to the broader loop between the first and second transition metal elements, beginning at period six. This loop of 14 elements is found at the bottom of the conventional table and often linked by thin lines to the place where this block of elements interrupts the table in the transition series. Often, each of the periods in this block of elements is named separately as the lanthanide and actinide series, as a pointer to where they come in the sequence of elements. Together this block of elements is also known as the f-block transition metals.

(b) The representative elements are shown in the “circular” section of the spiral periodic table and on the cylindrical section of the three-dimensional table. The d-block transition metals are shown in the “squarish” extension toward the lower right of the spiral table and on the broader loop of the three-dimensional table. The f-block transition metals are shown in the “rectangular” extension toward the upper right of the spiral table and on the thinner loop of the three-dimensional table. Note that both the spiral and three-dimensional tables represent the elements in a continuous series. In the conventional table (even the long form shown in the Web Companion interactive periodic table under “Electron Configuration”), there is a “break” at the end of each period which signals the completion of one of our shells, but does not signal how the elements simply continue in the same pattern in the next period.

Problem 4.70.
(a) The Schrödinger wave equation relates the wave properties of the electron to its kinetic energy. FALSE. The Schrödinger wave equation relates the wave properties of the electron to its total energy, which is a sum of kinetic and potential energy.

(b) The wave function \( \psi \) is a mathematical description of an allowed energy state (orbital) for an electron. TRUE.

(c) The wave function \( \psi \) is proportional to the probability of finding the electron in a given location near the nucleus. FALSE. The square of wave function \( \psi^2 \) is proportional to probability of finding the electron in given location around the nucleus.

(d) Orbitals with the same energy are called degenerate. TRUE.

(e) Orbitals can be thought of as a boxes into which electrons can be placed. FALSE. Orbitals describe calculated energy states of an electron and its corresponding spatial distribution. Orbitals do not exist (except as mathematical constructs), if there is no electron to be described.

(f) Each calculated energy level corresponds to a spatial probability distribution for the electron. TRUE. Keep in mind that each of the energy levels represented by a horizontal line on an
energy diagram for an atomic or molecular system describes a different spatial probability distribution for the electron wave. For a degenerate set of orbitals, each of the individual energy levels represents a different orbital shape and/or orientation, as in Figure 4.36 for the n = 2 orbitals.

Problem 4.71.
(a) This electron configuration for Al, 1s^2 2s^2 2p^4 3s^2 3p^3, has the correct number of electrons, 13, but the lower energy 2p orbital is not completely filled, as it should be for the lowest energy configuration. The ground-state configuration for Al is 1s^2 2s^2 2p^6 3s^2 3p^1.
(b) This electron configuration for P, 1s^2 2s^2 2p^6 3s^2 3p^3, represents only 14 electrons, but P has 15 electrons. The ground-state configuration for P is 1s^2 2s^2 2p^6 3s^2 3p^3.
(c) This electron configuration for B, 1s^2 2s^0 2p^3, has the correct number of electrons, 5, but the lower energy 2s orbital is empty, so this is not the lowest energy configuration. The ground-state configuration for B is 1s^2 2s^2 2p^1.

Problem 4.72.
The number of valence electrons located in an s orbital in each of these elements is (remember that the number of electrons in any s orbital is limited to two):
(a) potassium 1
(b) fluorine 2
(c) magnesium 2
(d) boron 2

Problem 4.73.
The number of valence electrons located in a p orbital in each of these elements is (remember that the number of electrons in any set of three degenerate p orbitals is limited to six):
(a) phosphorus 3
(b) aluminum 1
(c) lithium 0
(d) bromine 5
(e) sulfur 4

Problem 4.74.
Consider atoms A, B, and C with these electron configurations:

\[ A = 1s^2 2s^2 2p^6 3s^2 \quad B = 1s^2 2s^2 2p^6 3s^2 3p^4 \quad C = 1s^2 2s^2 2p^6 3s^2 3p^6 \]

Atom A will have the largest third ionization energy. In all three atoms, the first two electrons will come from the third shell and the first and second ionization energies will be in the order: C > B > A, because ionization energies increase across a period. However, the third electron in A must come from the second electron shell, which is closer to the nucleus than the third, so its electrons are much more tightly held. Thus, a good deal more energy will be required to remove the third electron from A than from either B or C, in which the third electron still comes from the third shell.
Problem 4.75.
Electron energy diagrams for the sulfur atom (16 electrons) and the sulfide ion (18 electrons), with Figure 4.37 as a template and Worked Example 4.60 as a guide, are:

![Energy Diagram for Sulfur Atom](image1)

![Energy Diagram for Sulfide Ion](image2)

Problem 4.76.
Electron energy diagrams for the lithium (3 electrons) and sodium (11 electrons) atoms, with Figure 4.37 as a template and Worked Example 4.60 as a guide, are:

![Energy Diagram for Lithium Atom](image3)

![Energy Diagram for Sodium Atom](image4)

Problem 4.77.
(a) The commonality of the series of atoms from H through Ar with the series of monocations He⁺ through K⁺ and the series of dications Li²⁺ through Ca²⁺ is that each series has the same pattern of electron configurations. That is, the electron configuration for H, He⁺, and Li²⁺ is 1s¹, for He, Li⁺, and Be²⁺, it is 1s², and so on up through Ar, K⁺, and Ca²⁺ with an electron configuration 1s²2s²2p⁶3s²3p⁶.

(b) We found that the patterns of first, second, and third ionization energies of elemental gaseous atoms as a function of atomic number (solutions to Problems 4.5 and 4.66) are very similar, except displaced to one higher atomic number as we go from first to second to third ionization energy. If the electron configurations of the atom and ions losing electrons are the same, then it makes sense that the patterns of next electron loss should also be the same.

Problem 4.78.
(a) Figure 4.37 showed the relative energy levels only up through the 4s orbital energy. The energy level scheme in the Web Companion interactive periodic table for “Electron Configuration” (reproduced here) carries the energy levels up through 7p. Energy increases
from the bottom to the top of the scheme. The $d$ (5-fold degenerate) and $f$ (7-fold degenerate) orbital energies (see Consider This 4.59) are included and can help you see how the periodic table is built up with its $s$, $p$, $d$, and $f$ blocks.

(b) The electron configuration for Li that you get from the Web Companion (shown above) is exactly what we wrote in Worked Example 4.60, with two $1s$ spin-paired electrons and one $2s$ electron.

(c) Our results in Consider This 4.62(a) give electron configurations for the elements Na through Ar that are exactly analogous to those for Li through Ne in Table 4.4 with the electron energies corresponding to $3s$ and $3p$ orbitals instead of $2s$ and $2p$. These are also the configurations you get from the Web Companion, as shown here for the Na and Ar configurations.

Problem 4.79.
(a) As you use the Web Companion interactive periodic table to examine the electron configurations of the first series of transition metals, Sc through Zn, you find that the pattern is not quite regular. The electron configurations for V, Cr, and Mn exemplify this irregularity:
Given the patterns for Sc, Ti, and V, our expectation would be that the electron configuration for Cr would be $[\text{Ar}]^{18}4s^23d^4$, instead of what is observed, $[\text{Ar}]^{18}4s^13d^5$. If the $4s$ energy levels were much lower in the energy than the $3d$, we would not expect this sort of irregularity. However, if the levels are rather close in energy, then the Cr atom might be lower in energy if the six highest energy electrons are all described by different orbitals. This arrangement spreads the electrons out a bit more so that electron-electron repulsion is a little less. The evidence suggests that the $4s$ and $3d$ electrons are of about the same energy.

(b) Note, as you found for simpler atoms in Consider This 4.62(c) and (d), that the electrons in the singly-occupied orbitals in these transition metals, all have the same spin. This arrangement of spins also serves to keep the electrons a bit farther apart on the average, further helps to reduce electron-electron repulsion, and hence helps stabilize the atom. It appears that, when electrons in an atom are in singly-occupied orbitals, the lowest energy state is when all the unpaired electrons have the same spin. This observation is known as Hund’s rule.

c) The electron configurations of the elements from Y through Cd do not follow exactly the same pattern as those from Sc through Zn. It appears that the $5s$ and $4d$ energy levels must be even closer together than the $4s$ and $3d$ levels, since electrons in this series often go preferentially to the $4d$ rather than $5s$ levels, as in Nb, Ru, Rh, and Pd.

**Problem 4.80.**

(a) The inset in Figure 4.27 shows that $E$ is a minimum at $R = 2$.

(b) Using the value of $R$ you from part (a) in equations (4.29) and (4.30) and assuming that the proportionality constant is the same in both cases, $H$, we find these potential and kinetic energies:

\[
PE = \frac{H}{R}; \quad \text{when } E \text{ is a minimum, } PE = \frac{H}{2}
\]

\[
KE = \frac{H}{R^2}; \quad \text{when } E \text{ is a minimum, } KE = \frac{H}{4}
\]

We see that the relationship of the kinetic energy to the potential energy of the proton-electron wave system is:

\[
KE = -\frac{1}{2} PE.
\]

(c) The total energy at its minimum for the proton-electron wave system as a function of $H$ from part (b) is:
\[ E = KE + PE = \frac{H}{4} - \frac{H}{2} = \frac{H}{4} \]

The ionization energy for the hydrogen atom, \( IE = 1312 \text{ kJ} \cdot \text{mol}^{-1} \), is the energy required to overcome the overall binding energy, \( E \), of the proton-electron system. Using this information, we can find the value of the proportionality constant, \( H \):

\[ IE = 1312 \text{ kJ} \cdot \text{mol}^{-1} = -E = \frac{H}{4} ; \quad H = 5248 \text{ kJ} \cdot \text{mol}^{-1} \]

**Problem 4.81.**

(a) For an electron (charge \(-1\)) attracted by a nucleus of charge \(+Z\), the potential energy is:

\[ PE = \frac{Z}{r} \]

(b) By analogy with the transformation of equation (4.16) to equation (4.30), we can write the potential energy for a spherical electron wave attracted to the nucleus in part (a), using the proportionality constant, \( H \), from Problem 4.80(b):

\[ PE = \frac{ZH}{R} \]

(c) If the kinetic energy of the electron wave is independent of the nuclear charge, the total energy of a one-electron ion with a nuclear charge \(+Z\), as a function of \( R \), \( H \), and \( Z \), is:

\[ E = KE + PE = \frac{H}{R^2} - \frac{ZH}{R} \]

(d) Three methods for finding the value of \( R \) that makes the total energy in part (c) a minimum are:

- Method (i), using calculus, gives this result:

\[ \frac{dE}{dR} = 0 = \frac{2H}{R^3} + \frac{ZH}{R^2} = \frac{2H}{R^3} \cdot \frac{ZH}{R^2} \]

\[ 2H = ZHR \quad R = \frac{2}{Z} \]

- Method (ii), for \( Z = 1, 2, 3, \) and \( 5 \), gives \( R = 2 \) (from Figure 4.27), 1, 0.67, and 0.40 from the plot below. (Substituting these values of \( Z \) into the result from method (i) gives the same \( R \) values we get from the plot, so the methods produce the same results.)
• Method (iii), assuming $KE = -PE/2$ gives:

$$\frac{H}{R^2} = \frac{ZH}{2R}$$

Solving for $R$ gives the same result as in method (i).

(e) Substitute the result from part (d) into the expression for $E$ from part (c) to get the minimum for the total energy of a one-electron ion as a function of $H$ and $Z$:

$$E = \frac{H}{2/Z^2} - \frac{ZH}{2/Z} = \frac{Z^2H}{4} - \frac{Z^2H}{2} = -\frac{Z^2H}{4}$$

For He$^+$, Li$^{2+}$, and B$^{4+}$, the $Z$ values are 2, 3, and 5, so the total energies are $H$, $(9/4)H$, and $(25/4)H$. From Problem 4.80(c), we know that $H = 5248 \text{ kJ} \cdot \text{mol}^{-1}$, so the ionization energies ($-E$) are 5248, 11,808, and 32,800 kJ·mol$^{-1}$, respectively. As you see in Table 4.2, these are exactly the ionization energies for the final electron in these three atoms. You can calculate the last ionization energy for any atom. You can check your results for the final ionization energy for any of the atoms through Ne in the table. (The last few values in the table are only given to three significant figures.)

**Problem 4.82.**

[NOTE: The sources for this response are the Department of Physics and Astronomy, Arizona State University, Tempe, AZ (accept.la.asu.edu/PiN/rdg/elmicr/versus.shtml) and Allen R. Sampson, Advanced Research Systems (www.sem.com/analytic/sem.htm). The on-line materials have been extensively edited to include only what is necessary to see how the information in Chapter 4 is used to respond to Problem 4.82.]

The attractive feature of optical microscopy is that it is so easy: samples can be analyzed in air or water, the images are in natural color with magnifications of up to one hundred to one thousand times, and modern semiconductor electronics with charge-coupled devices allow image processing. The optical microscope should dominate the field of microscopy, but it doesn't. The scanning electron microscope (SEM) is the microscope of choice because of its depth of focus and resolving capability. Examination of this figure shows a striking contrast between an optical and SEM image of a radiolarian (tiny marine creature) at the same magnification. [Figure taken from J.I. Goldstein *et al.*., eds., *Scanning Electron Microscopy and X-Ray Microanalysis*, (Plenum Press, New York, 1980).]
In the optical micrograph taken at high resolution only a section of the radiolarian is in sharp focus. In the SEM image the whole specimen is in focus. The three-dimensional appearance of this image is a direct result of the large depth of field of the SEM. It is this large depth of field in the SEM images that is one of its most attractive features.

The other important feature of the SEM is resolving power, which is the smallest detail that a microscope can resolve, or "see". Electron microscopy takes advantage of the wave nature of rapidly moving electrons. The resolving power of electron microscopes is orders of magnitude better than that of an optical microscope because the wavelength of the probing beam is orders of magnitude smaller. The limit on what size can be resolved—regardless of instrument type—is set by the wavelength. The wavelength of the visible light used in optical microscopes is between 400 and 700 nm. The resolving power of high-quality light microscopes is limited by diffraction of the imaging light to about 200 nm (0.2 \( \mu \)m). Scanning electron microscopy uses electrons with energies of a few thousand electron volts (eV), energies a thousand times greater than that of visible light (2 to 3 eV). The wavelength of the electron is given by \( \frac{h}{m \cdot u} \). For 3600 eV electrons, the wavelength is 0.02 nm (20 pm). The resolving power of ordinary electron microscopes is about 1 nm (not as good as the theoretical value because construction details limit resolving power), a value that can be pushed to 0.1 nanometer.

**Problem 4.83.**

Einstein’s relativity theory shows that the effective mass of a moving particle, \( m_{\text{moving}} \), is related to its rest mass, \( m_{\text{rest}} \), its velocity, \( u \), and the speed of light, \( c \), by this equation:

\[
m_{\text{moving}} = \frac{m_{\text{rest}}}{\sqrt{1 - \frac{u}{c}^2}}
\]

(a) The "relativistic mass" of a 145 g baseball thrown at 44 m·s\(^{-1}\) is:

\[
m_{\text{moving}} = \frac{145 \text{ g}}{\sqrt{1 - \frac{44 \text{ m} \cdot \text{s}^{-1}}{3.0 \times 10^8 \text{ m} \cdot \text{s}^{-1}}^2}} = \frac{145 \text{ g}}{\sqrt{1 - 2.2 \times 10^{-14}}} = \frac{145 \text{ g}}{1} = 145 \text{ g}
\]

At such a slow rate of travel, the baseball is not affected by relativity and batters do not have to be concerned about learning relativity theory (at least in order to play baseball).

(b) For the alpha particle in Problem 4.3(b) moving at 1.5 \( \times 10^7 \text{ m} \cdot \text{s}^{-1} \) we have an effective mass of:
\[ m_{\text{moving}} = \frac{6.6 \times 10^{-24} \text{ g}}{\sqrt{1.5 \times 10^7 \text{ m s}^{-1} \div 3.0 \times 10^8 \text{ m s}^{-1}}} = \frac{6.6 \times 10^{-24} \text{ g}}{\sqrt{0.998}} = \frac{6.6 \times 10^{-24} \text{ g}}{0.999} = 6.6 \times 10^{-24} \text{ g} \]

Within the number of significant digits in the data, this velocity does not affect the mass of the alpha particle and the calculations in Problem 4.3(b) do not need correction.

(c) For an electron moving at 50% the speed of light, the \( u/c \) ratio is 0.50, so we have for the effective mass of the electron:

\[ m_{\text{moving}} = \frac{9.1 \times 10^{-31} \text{ kg}}{\sqrt{0.50^2}} = \frac{9.1 \times 10^{-31} \text{ kg}}{0.75} = \frac{9.1 \times 10^{-31} \text{ kg}}{0.87} = 10.5 \times 10^{-31} \text{ kg} \]

The effective mass of the moving electron is about 14% higher than the rest mass. Substitution into the de Broglie wavelength equation gives:

\[ \frac{h}{m u} = \frac{6.63 \times 10^{-34} \text{ J s}}{10.5 \times 10^{-31} \text{ kg} \cdot 1.50 \times 10^8 \text{ m s}^{-1}} = 4.2 \times 10^{-12} \text{ m} = 4.2 \text{ pm} \]

This electron has a wavelength comparable to gamma radiation.

Problem 4.84.

(a) This problem, finding the wavelength of a proton that has been accelerated to 50% the speed of light, is identical to Problem 4.83(c), except that the rest mass of the particle in this case is \( 1.67 \times 10^{-27} \text{ kg} \):

\[ m_{\text{moving}} = \frac{1.67 \times 10^{-27} \text{ kg}}{0.87} = 1.92 \times 10^{-27} \text{ kg} \]

\[ m_{\text{moving}} = \frac{6.63 \times 10^{-34} \text{ J s}}{10.5 \times 10^{-31} \text{ kg} \cdot 1.50 \times 10^8 \text{ m s}^{-1}} = 2.3 \times 10^{-15} \text{ m} = 2.3 \text{ fm (femtometer)} \]

(b) Let’s do this problem in reverse. First we will calculate the kinetic energy of one electron moving at 90% the speed of light (\( u/c = 0.90 \)) and then multiply by Avogadro’s number to get the energy of a mole of these electrons.

\[ m_{\text{moving}} = \frac{9.1 \times 10^{-31} \text{ kg}}{\sqrt{0.90^2}} = \frac{9.1 \times 10^{-31} \text{ kg}}{0.44} = 20.9 \times 10^{-31} \text{ kg} \]

The kinetic energy of this electron is:

\[ KE \text{ (one electron)} = \frac{m_{\text{moving}} u^2}{2} = \frac{20.9 \times 10^{-31} \text{ kg} \cdot 0.90 \cdot 3.00 \times 10^8 \text{ m s}^{-1}^2}{2} = 7.6 \times 10^{-14} \text{ J} = 7.6 \times 10^{-11} \text{ kJ} \]

A mole of these electrons has:

\[ KE \text{ (mole)} = (7.6 \times 10^{-11} \text{ kJ}) (6.02 \times 10^{23} \text{ mol}^{-1}) = 4.6 \times 10^{10} \text{ kJ mol}^{-1} \]

This is enough energy to raise the temperature of 11 million liters of water from 0 °C to 100 °C.

Problem 4.85.

(a) The representation of the periodic table shown in the “Electron Configuration” module in the Web Companion interactive periodic table looks like this (without the block designations):
The version of the periodic table on the inside front cover of this book is the one you will probably see most often, because it conserves space and can easily be printed large enough to be easily readable on a standard size page. Compared to the form from the Companion shown here, we see that the f-block elements (often called the inner transition series or the lanthanides and actinides on the book cover) have been moved below the rest of the table and the s- and d-blocks moved together to make the table less long (wide). Using the form from the book cover, we have to be careful to take note of the jump in atomic number (and hence properties) between elements 56 (barium) and 71 (lutetium) and between elements 88 (radium) and 103 (lawrencium) where the inner transition series elements have been shifted out.

(b) In the spiral periodic table, Figure 4.34, the s- and p-block elements, usually called the representative elements, together form the “circular” (actually a spiral) portion on the left. The d-block elements are represented by the broader rectangular extension of the spiral on the right and the f-block elements by the narrower rectangular extension. The most interesting difference between the spiral table and the one above is that the spiral table is continuous, that is, the noble gas elements are flanked by a halogen and an alkali metal, rather than ending a period. The families of elements (periodicity) are represented by elements going out on a radius (or stacked in successive layers), but the periods themselves are a little more difficult to discern.

(c) The three-dimensional table shown in Problem 4.69 and the chapter opening illustration is related to the one above much as described for the spiral table in part (b). The cylinder that is the tallest part of the three-dimensional table is a helix (spiral), much like the circular spiral in the spiral table and also consists of the s- and p-block (representative) elements. Each cycle around the cylinder represents a period of the table and the families are in the columns down the surface of cylinder. The loop extending directly from the cylinder are the d-block transition elements and “interrupt” the spiral, effectively making each new loop of spiral longer, as they make the table above longer (wider). Another loop, the f-block elements, extends from the d-block element loop, which, again, effectively lengthens the spiral.

[NOTE: There is a somewhat difference between the ways the “linear” tables inside the cover of the text and in part (a) and the spiral tables described here and in part (b) treat the initial elements of the f- and d-blocks. Note that the linear tables place Lu and Lr in the family that includes the first elements in the d-block, whereas, in the spiral tables, La and Ac are in this position. The argument for putting La and Ac in this position is that these elements have no f electrons in their valence shells and their electron configurations are (n-1)dns2, exactly the same as the first elements in the d-block (periods 4 and 5). The electron configurations of Lu and Lr are (n-2)fn4(n-1)dns2, which are also like the electron configurations of the first elements in the d-block, if the filled inner shell of f electrons are assumed to be core electrons not valence electrons. The predominant chemistry for all four elements, formation of 3+ metal cations, is the same as that for the 4th and 5th period elements in this family, so it is difficult to differentiate between these conventions on that basis.]