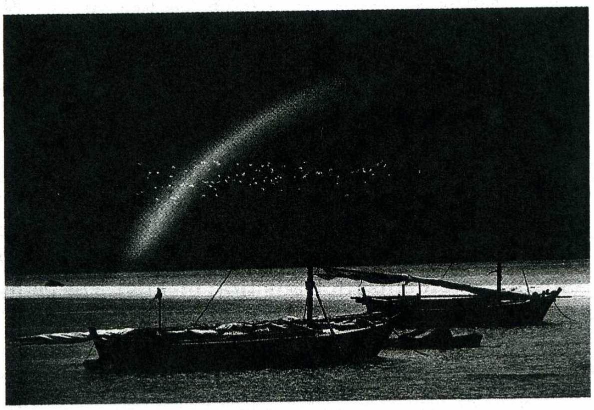
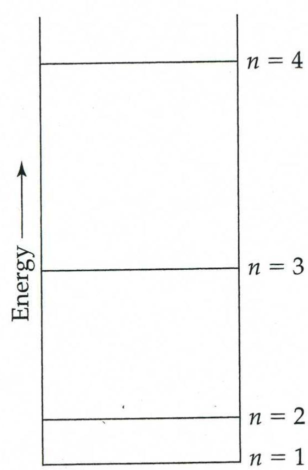
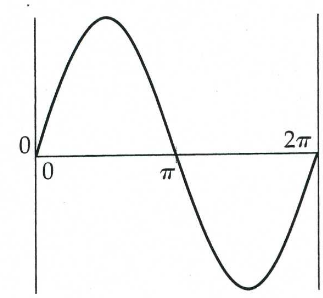
6.4 The familiar phenomenon of a rainbow results from the
CQ diffraction of sunlight through raindrops. **(a)** Does the wavelength of light increase or decrease as we proceed outward from the innermost band of the rainbow? **(b)** Does the frequency of light increase or decrease as we proceed outward? **(c)** Suppose that instead of sunlight, the visible light from a hydrogen discharge tube (Figure 6.11) was used as the light source. What do you think the resulting "hydrogen discharge rainbow" would look like? [Section 6.3]



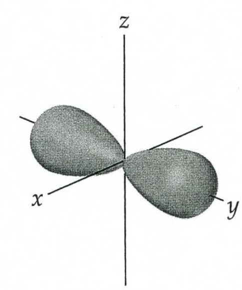
6.5 A certain quantum mechanical system has the energy
CQ levels shown in the diagram below. The energy levels are indexed by a single quantum number n that is an integer. **(a)** As drawn, which quantum numbers are involved in the transition that requires the most energy? **(b)** Which quantum numbers are involved in the transition that requires the least energy? **(c)** Based on the drawing, put the following in order of increasing wavelength of the light absorbed or emitted during the transition: (*i*) $n = 1$ to $n = 2$; (*ii*) $n = 3$ to $n = 2$; (*iii*) $n = 2$ to $n = 4$; (*iv*) $n = 3$ to $n = 1$. [Section 6.3]



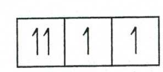
6.6 Consider a fictitious one-dimensional system with on
CQ electron. The wave function for the electron, draw below, is $\psi(x) = \sin x$ from $x = 0$ to $x = 2\pi$. **(a)** Sketc the probability density, $\psi^2(x)$, from $x = 0$ to $x = 2\pi$ **(b)** At what value or values of x will there be the greates probability of finding the electron? **(c)** What is the prob ability that the electron will be found at $x = \pi$? What i such a point in a wave function called? [Section 6.5]



6.7 The contour representation of one of the orbitals for th
CQ $n = 3$ shell of a hydrogen atom is shown belov **(a)** What is the quantum number l for this orbita **(b)** How do we label this orbital? **(c)** How would yo modify this sketch to show the analogous orbital for th $n = 4$ shell? [Section 6.6]



6.8 The drawing below shows part of the orbital diagra
CQ for an element. **(a)** As drawn, the drawing is *incorrec* Why? **(b)** How would you correct the drawing withou changing the number of electrons? **(c)** To which group the periodic table does the element belong? [Section 6.



EXERCISES

Radiant Energy

6.9 What are the basic SI units for **(a)** the wavelength of light, **(b)** the frequency of light, **(c)** the speed of light?

6.10 **(a)** What is the relationship between the wavelength and the frequency of radiant energy? **(b)** Ozone in the upper atmosphere absorbs energy in the 210–230-n range of the spectrum. In what region of the electroma netic spectrum does this radiation occur?

6.11 Label each of the following statements as true or fals
CQ For those that are false, correct the statement. **(a)** Visib

light is a form of electromagnetic radiation. **(b)** The frequency of radiation increases as the wavelength increases. **(c)** Ultraviolet light has longer wavelengths than visible light. **(d)** Electromagnetic radiation and sound waves travel at the same speed.

6.12 **CQ** Determine which of the following statements are false, and correct them. **(a)** Electromagnetic radiation is incapable of passing through water. **(b)** Electromagnetic radiation travels through a vacuum at a constant speed, regardless of wavelength. **(c)** Infrared light has higher frequencies than visible light. **(d)** The glow from a fireplace, the energy within a microwave oven, and a foghorn blast are all forms of electromagnetic radiation.

6.13 **CQ** Arrange the following kinds of electromagnetic radiation in order of increasing wavelength: infrared, green light, red light, radio waves, X rays, ultraviolet light.

6.14 **CQ** List the following types of electromagnetic radiation in order of increasing wavelength: **(a)** the gamma rays produced by a radioactive nuclide used in medical imaging; **(b)** radiation from an FM radio station at 93.1 MHz on the dial; **(c)** a radio signal from an AM radio station at 680 kHz on the dial; **(d)** the yellow light from sodium vapor streetlights; **(e)** the red light of a light-emitting diode, such as in a calculator display.

6.15 **(a)** What is the frequency of radiation that has a wavelength of 955 μm? **(b)** What is the wavelength of radiation that has a frequency of $5.50 \times 10^{14}\ s^{-1}$? **(c)** Would the radiations in part (a) or part (b) be visible to the human eye? **(d)** What distance does electromagnetic radiation travel in 50.0 μs?

6.16 **(a)** What is the frequency of radiation whose wavelength is 10.0 Å? **(b)** What is the wavelength of radiation that has a frequency of $7.6 \times 10^{10}\ s^{-1}$? **(c)** Would the radiations in part (a) or part (b) be detected by an X-ray detector? **(d)** What distance does electromagnetic radiation travel in 25.5 fs?

6.17 Excited mercury atoms emit light strongly at a wavelength of 436 nm. What is the frequency of this radiation? Using Figure 6.4, predict the color associated with this wavelength.

6.18 An argon ion laser emits light at 489 nm. What is the frequency of this radiation? Is this emission in the visible spectrum? If yes, what color is it?

Quantized Energy and Photons

6.19 **CQ** **(a)** What does it mean when we say energy is quantized? **(b)** Why don't we notice the quantization of energy in everyday activities?

6.20 **CQ** Einstein's 1905 paper on the photoelectric effect was the first important application of Planck's quantum hypothesis. Describe Planck's original hypothesis, and explain how Einstein made use of it in his theory of the photoelectric effect.

6.21 **(a)** Calculate the smallest increment of energy (a quantum) that can be emitted or absorbed at a wavelength of 438 nm. **(b)** Calculate the energy of a photon of frequency $6.75 \times 10^{12}\ s^{-1}$. **(c)** What wavelength of radiation has photons of energy 2.87×10^{-18} J? In what portion of the electromagnetic spectrum would this radiation be found?

6.22 **(a)** Calculate the smallest increment of energy that can be emitted or absorbed at a wavelength of 10.8 mm. **(b)** Calculate the energy of a photon from an FM radio station at a frequency of 101.1 MHz. **(c)** For what frequency of radiation will a mole of photons have energy 24.7 kJ? In what region of the electromagnetic spectrum would this radiation be found?

6.23 **(a)** Calculate and compare the energy of a photon of wavelength 3.3 μm with that of wavelength 0.154 nm. **(b)** Use Figure 6.4 to identify the region of the electromagnetic spectrum to which each belongs.

6.24 An AM radio station broadcasts at 1010 kHz, and its FM partner broadcasts at 98.3 MHz. Calculate and compare the energy of the photons emitted by these two radio stations.

6.25 One type of sunburn occurs on exposure to UV light of wavelength in the vicinity of 325 nm. **(a)** What is the energy of a photon of this wavelength? **(b)** What is the energy of a mole of these photons? **(c)** How many photons are in a 1.00 mJ burst of this radiation?

6.26 The energy from radiation can be used to cause the rupture of chemical bonds. A minimum energy of 941 kJ/mol is required to break the nitrogen-nitrogen bond in N_2. What is the longest wavelength of radiation that possesses the necessary energy to break the bond? What type of electromagnetic radiation is this?

6.27 A diode laser emits at a wavelength of 987 nm. **(a)** In what portion of the electromagnetic spectrum is this radiation found? **(b)** All of its output energy is absorbed in a detector that measures a total energy of 0.52 J over a period of 32 s. How many photons per second are being emitted by the laser?

6.28 A stellar object is emitting radiation at 3.55 mm. **(a)** What type of electromagnetic spectrum is this radiation? **(b)** If the detector is capturing 3.2×10^8 photons per second at this wavelength, what is the total energy of the photons detected in one hour?

6.29 Molybdenum metal must absorb radiation with a minimum frequency of $1.09 \times 10^{15}\ s^{-1}$ before it can emit an electron from its surface via the photoelectric effect. **(a)** What is the minimum energy needed to produce this effect? **(b)** What wavelength radiation will provide a photon of this energy? **(c)** If molybdenum is irradiated with light of wavelength of 120 nm, what is the maximum possible kinetic energy of the emitted electrons?

6.30 It requires a photon with a minimum energy of 4.41×10^{-19} J to emit electrons from sodium metal. **(a)** What is the minimum frequency of light necessary to emit electrons from sodium via the photoelectric effect? **(b)** What is the wavelength of this light? **(c)** If sodium is irradiated with light of 439 nm, what is the maximum possible kinetic energy of the emitted electrons? **(d)** What is the maximum number of electrons that can be freed by a burst of light whose total energy is 1.00 μJ?

Bohr's Model; Matter Waves

6.31 Explain how the existence of line spectra is consistent **CQ** with Bohr's theory of quantized energies for the electron in the hydrogen atom.

6.32 **(a)** In terms of the Bohr theory of the hydrogen atom, what **CQ** process is occurring when excited hydrogen atoms emit radiant energy of certain wavelengths and only those wavelengths? **(b)** Does a hydrogen atom "expand" or "contract" as it moves from its ground state to an excited state?

6.33 Is energy emitted or absorbed when the following elec- **CQ** tronic transitions occur in hydrogen? **(a)** from $n = 4$ to $n = 2$, **(b)** from an orbit of radius 2.12 Å to one of radius 8.46 Å, **(c)** an electron adds to the H^+ ion and ends up in the $n = 3$ shell.

6.34 Indicate whether energy is emitted or absorbed when **CQ** the following electronic transitions occur in hydrogen: **(a)** from $n = 2$ to $n = 6$, **(b)** from an orbit of radius 4.76 Å to one of radius 0.529 Å, **(c)** from the $n = 6$ to the $n = 9$ state.

6.35 Using Equation 6.5, calculate the energy of an electron in the hydrogen atom when $n = 2$ and when $n = 6$. Calculate the wavelength of the radiation released when an electron moves from $n = 6$ to $n = 2$. Is this line in the visible region of the electromagnetic spectrum? If so, what color is it?

6.36 For each of the following electronic transitions in the hydrogen atom, calculate the energy, frequency, and wavelength of the associated radiation, and determine whether the radiation is emitted or absorbed during the transition: **(a)** from $n = 4$ to $n = 1$, **(b)** from $n = 5$ to $n = 2$, **(c)** from $n = 3$ to $n = 6$. Does any of these transitions emit or absorb visible light?

6.37 The visible emission lines observed by Balmer all involved $n_f = 2$. **(a)** Explain why only the lines with $n_f = 2$ were observed in the visible region of the electromagnetic spectrum. **(b)** Calculate the wavelengths of the first three lines in the Balmer series—those for which $n_i = 3, 4,$ and 5—and identify these lines in the emission spectrum shown in Figure 6.12.

6.38 The Lyman series of emission lines of the hydrogen atom are those for which $n_f = 1$. **(a)** Determine the region of the electromagnetic spectrum in which the lines of the Lyman series are observed. **(b)** Calculate the wavelengths of the first three lines in the Lyman series—those for which $n_i = 2, 3,$ and 4.

[6.39] One of the emission lines of the hydrogen atom has a wavelength of 93.8 nm. **(a)** In what region of the electromagnetic spectrum is this emission found? **(b)** Determine the initial and final values of n associated with this emission.

[6.40] The hydrogen atom can absorb light of wavelength 2626 nm. **(a)** In what region of the electromagnetic spectrum is this absorption found? **(b)** Determine the initial and final values of n associated with this absorption.

6.41 Use the de Broglie relationship to determine the wavelengths of the following objects: **(a)** an 85-kg person skiing at 50 km/hr, **(b)** a 10.0-g bullet fired at 250 m/s, **(c)** a lithium atom moving at 2.5×10^5 m/s.

6.42 Among the elementary subatomic particles of physics is the muon, which decays within a few nanoseconds after formation. The muon has a rest mass 206.8 times that of an electron. Calculate the de Broglie wavelength associated with a muon traveling at a velocity of 8.85×10^5 cm/s.

6.43 Neutron diffraction is an important technique for determining the structures of molecules. Calculate the velocity of a neutron that has a characteristic wavelength of 0.955 Å. (Refer to the inside cover for the mass of the neutron.)

6.44 The electron microscope has been widely used to obtain highly magnified images of biological and other types of materials. When an electron is accelerated through a particular potential field, it attains a speed of 9.38×10^6 m/s. What is the characteristic wavelength of this electron? Is the wavelength comparable to the size of atoms?

6.45 Using Heisenberg's uncertainty principle, calculate the uncertainty in the position of **(a)** a 1.50-mg mosquito moving at a speed of 1.40 m/s if the speed is known to within ± 0.01 m/s; **(b)** a proton moving at a speed of $(5.00 \pm 0.01) \times 10^4$ m/s. (The mass of a proton is given in the table of fundamental constants in the inside cover of the text.)

6.46 Calculate the uncertainty in the position of **(a)** an electron moving at a speed of $(3.00 \pm 0.01) \times 10^5$ m/s, **(b)** a neutron moving at this same speed. (The masses of an electron and a neutron are given in the table of fundamental constants in the inside cover of the text.) **(c)** What are the implications of these calculations to our model of the atom?

Quantum Mechanics and Atomic Orbitals

6.47 **(a)** Why does the Bohr model of the hydrogen atom vio- **CQ** late the uncertainty principle? **(b)** In what way is the description of the electron using a wave function consistent with de Broglie's hypothesis? **(c)** What is meant by the term *probability density*? Given the wave function, how do we find the probability density at a certain point in space?

6.48 **(a)** According to the Bohr model, an electron in the **CQ** ground state of a hydrogen atom orbits the nucleus at a specific radius of 0.53 Å. In the quantum mechanical description of the hydrogen atom, the most probable distance of the electron from the nucleus is 0.53 Å. Why are these two statements different? **(b)** Why is the use of Schrödinger's wave equation to describe the location of a particle very different from the description obtained from classical physics? **(c)** In the quantum mechanical description of an electron, what is the physical significance of the square of the wave function, ψ^2?

6.49 **(a)** For $n = 4$, what are the possible values of l? **(b)** For $l = 2$, what are the possible values of m_l?

6.50 How many possible values for l and m_l are there when **(a)** $n = 3$; **(b)** $n = 5$?

6.51 Give the numerical values of n and l corresponding to each of the following designations: **(a)** $3p$, **(b)** $2s$, **(c)** $4f$, **(d)** $5d$.

6.52 Give the values for n, l, and m_l for **(a)** each orbital in the $2p$ subshell, **(b)** each orbital in the $5d$ subshell.

6.53 **CQ** Which of the following represent impossible combinations of n and l: **(a)** $1p$, **(b)** $4s$, **(c)** $5f$, **(d)** $2d$?

6.54 **CQ** Which of the following are permissible sets of quantum numbers for an electron in a hydrogen atom: **(a)** $n = 2$, $l = 1$, $m_l = 1$; **(b)** $n = 1$, $l = 0$, $m_l = -1$; **(c)** $n = 4$, $l = 2$, $m_l = -2$; **(d)** $n = 3$, $l = 3$, $m_l = 0$? For those combinations that are permissible, write the appropriate designation for the subshell to which the orbital belongs (that is, $1s$, and so on).

6.55 Sketch the shape and orientation of the following types of orbitals: **(a)** s, **(b)** p_z, **(c)** d_{xy}.

6.56 Sketch the shape and orientation of the following types of orbitals: **(a)** p_x, **(b)** d_{z^2}, **(c)** $d_{x^2-y^2}$.

6.57 **CQ** **(a)** What are the similarities and differences between the $1s$ and $2s$ orbitals of the hydrogen atom? **(b)** In what sense does a $2p$ orbital have directional character? Compare the "directional" characteristics of the p_x and $d_{x^2-y^2}$ orbitals (that is, in what direction or region of space is the electron density concentrated?). **(c)** What can you say about the average distance from the nucleus of an electron in a $2s$ orbital as compared with a $3s$ orbital? **(d)** For the hydrogen atom, list the following orbitals in order of increasing energy (that is, most stable ones first): $4f$, $6s$, $3d$, $1s$, $2p$.

6.58 **CQ** **(a)** With reference to Figure 6.18, what is the relationship between the number of nodes in an s orbital and the value of the principal quantum number? **(b)** Identify the number of nodes; that is, identify places where the electron density is zero, in the $2p_x$ orbital; in the $3s$ orbital. **(c)** What information is obtained from the radial probability functions in Figure 6.18? **(d)** For the hydrogen atom, list the following orbitals in order of increasing energy: $3s$, $2s$, $2p$, $5s$, $4d$.

Many-Electron Atoms and Electron Configurations

6.59 **CQ** For a given value of the principal quantum number, n, how do the energies of the s, p, d, and f subshells vary for **(a)** hydrogen, **(b)** a many-electron atom?

6.60 **CQ** **(a)** The average distance from the nucleus of a $3s$ electron in a chlorine atom is smaller than that for a $3p$ electron. In light of this fact, which orbital is higher in energy? **(b)** Would you expect it to require more or less energy to remove a $3s$ electron from the chlorine atom, as compared with a $2p$ electron? Explain.

6.61 **CQ** **(a)** What are the possible values of the electron spin quantum number? **(b)** What piece of experimental equipment can be used to distinguish electrons that have different values of the electron spin quantum number? **(c)** Two electrons in an atom both occupy the $1s$ orbital. What quantity must be different for the two electrons? What principle governs the answer to this question?

6.62 **CQ** **(a)** State the Pauli exclusion principle in your own words. **(b)** The Pauli exclusion principle is, in an important sense, the key to understanding the periodic table. Explain why.

6.63 What is the maximum number of electrons that can occupy each of the following subshells: **(a)** $3p$, **(b)** $5d$, **(c)** $2s$, **(d)** $4f$?

6.64 What is the maximum number of electrons in an atom that can have the following quantum numbers: **(a)** $n = 2$, $m_s = -\frac{1}{2}$; **(b)** $n = 5$, $l = 3$; **(c)** $n = 4$, $l = 3$, $m_l = -3$; **(d)** $n = 4$, $l = 1$, $m_l = 1$.

6.65 **CQ** **(a)** What does each box in an orbital diagram represent? **(b)** What quantity is represented by the direction (either up or down) of the half arrows in an orbital diagram? **(c)** Is Hund's rule needed to write the electron configuration of beryllium? Explain.

6.66 **CQ** **(a)** What are "valence electrons"? **(b)** What are "unpaired electrons"? **(c)** How many valence electrons does a P atom possess? How many of these are unpaired?

6.67 Write the condensed electron configurations for the following atoms, using the appropriate noble-gas core abbreviations: **(a)** Cs, **(b)** Ni, **(c)** Se, **(d)** Cd, **(e)** Ac, **(f)** Pb.

6.68 Write the condensed electron configurations for the following atoms, and indicate how many unpaired electrons each has: **(a)** Ga, **(b)** Ca, **(c)** V, **(d)** I, **(e)** Y, **(f)** Pt, **(g)** Lu.

6.69 Meitnerium, Mt, element 109, named after Lisa Meitner, is a transition metal expected to have the same outer-electron configuration as iridium. By using this observation (and without looking at Figure 6.30), write the electron configuration of meitnerium. Use [Rn] to represent the first 86 electrons of the electron configuration.

6.70 In 1999 it was reported that element 118 had been made artificially. This report was retracted in 2001, and since then there have been no other claims that element 118 can be made. **(a)** Write the electron configuration for element 118. **(b)** How many unpaired electrons would you expect an atom of element 118 to have? **(c)** To which group of elements in the periodic table would you expect element 118 to belong?

6.71 Identify the specific element that corresponds to each of the following electron configurations: **(a)** $1s^2 2s^2 2p^6 3s^2$, **(b)** $[Ne]3s^2 3p^1$, **(c)** $[Ar]4s^1 3d^5$, **(d)** $[Kr]5s^2 4d^{10} 5p^4$.

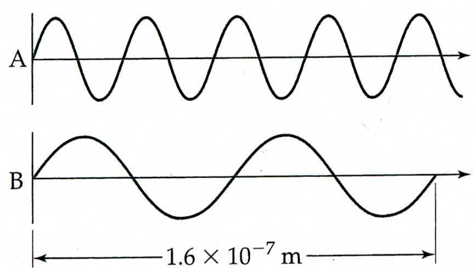
6.72 Identify the group of elements that corresponds to each of the following generalized electron configurations:
(a) [noble gas] $ns^2 np^5$
(b) [noble gas] $ns^2(n-1)d^2$
(c) [noble gas] $ns^2(n-1)d^{10}np^1$
(d) [noble gas] $ns^2(n-2)f^6$

6.73 **CQ** What is wrong with the following electron configurations for atoms in their ground states? **(a)** $1s^2 2s^2 3s^1$, **(b)** $[Ne]2s^2 2p^3$, **(c)** $[Ne]3s^2 3d^5$.

6.74 The following electron configurations represent excited states. Identify the element, and write its ground-state condensed electron configuration. **(a)** $1s^2 2s^2 3p^2 4p^1$, **(b)** $[Ar]3d^{10} 4s^1 4p^4 5s^1$, **(c)** $[Kr]4d^6 5s^2 5p^1$.

Additional Exercises

6.75 Consider the two waves shown here, which we will consider to represent two electromagnetic radiations:
(a) What is the wavelength of wave A? Of wave B?
(b) What is the frequency of wave A? Of wave B?
(c) Identify the regions of the electromagnetic spectrum to which waves A and B belong.



6.76 Certain elements emit light of a specific wavelength when they are burned. Historically, chemists used such emission wavelengths to determine whether specific elements were present in a sample. Some characteristic wavelengths for some of the elements are

Ag	328.1 nm	Fe	372.0 nm
Au	267.6 nm	K	404.7 nm
Ba	455.4 nm	Mg	285.2 nm
Ca	422.7 nm	Na	589.6 nm
Cu	324.8 nm	Ni	341.5 nm

(a) Determine which elements emit radiation in the visible part of the spectrum. **(b)** Which element emits photons of highest energy? Of lowest energy? **(c)** When burned, a sample of an unknown substance is found to emit light of frequency $6.59 \times 10^{14}\ s^{-1}$. Which of these elements is probably in the sample?

6.77 In June 2004, the Cassini–Huygens spacecraft began orbiting Saturn and transmitting images to Earth. The closest distance between Saturn and Earth is 746 million miles. What is the minimum amount of time it takes for the transmitted signals to travel from the spacecraft to Earth?

6.78 The rays of the Sun that cause tanning and burning are in the ultraviolet portion of the electromagnetic spectrum. These rays are categorized by wavelength: So-called UV-A radiation has wavelengths in the range of 320–380 nm, whereas UV-B radiation has wavelengths in the range of 290–320 nm. **(a)** Calculate the frequency of light that has a wavelength of 320 nm. **(b)** Calculate the energy of a mole of 320-nm photons. **(c)** Which are more energetic, photons of UV-A radiation or photons of UV-B radiation? **(d)** The UV-B radiation from the Sun is considered a greater cause of sunburn in humans than is UV-A radiation. Is this observation consistent with your answer to part (c)?

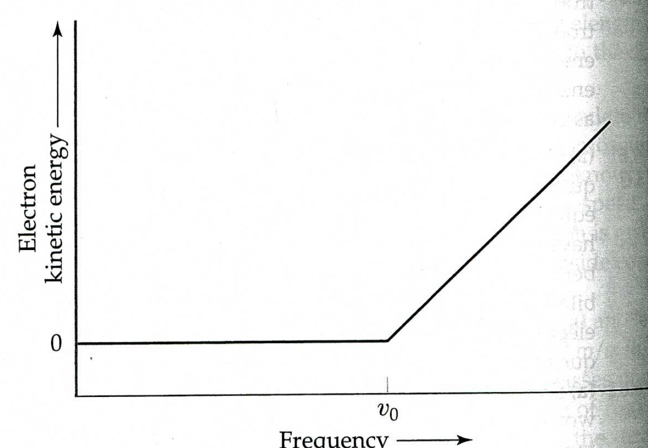
6.79 The watt is the derived SI unit of power, the measure of energy per unit time: 1 W = 1 J/s. A semiconductor laser in a CD player has an output wavelength of 780 nm and a power level of 0.10 mW. How many photons strike the CD surface during the playing of a CD 69 minutes in length?

6.80 Carotenoids, present in all organisms capable of photosynthesis, extend the range of light absorbed by the organism. They exhibit maximal capacity for absorption of light in the range of 440–470 nm. Calculate the energy in kilojoules represented by absorption of an Avogadro's number of photons of wavelength 455 nm.

[6.81] A photocell, such as the one illustrated in Figure 6.7(b), is a device used to measure the intensity of light. In a certain experiment, when light of wavelength 630 nm is directed onto the photocell, electrons are emitted at the rate of 2.6×10^{-12} C/s. Assume that each photon that impinges on the photocell emits one electron. How many photons per second are striking the photocell? How much energy per second is the photocell absorbing?

6.82 The light-sensitive substance in black-and-white photographic film is AgBr. Photons provide the energy necessary to transfer an electron from Br^- to Ag^+ to produce Ag and Br and thereby darken the film. **(a)** If a minimum energy of 2.00×10^5 J/mol is needed for this process, what is the minimum energy needed by each photon? **(b)** Calculate the wavelength of the light necessary to provide photons of this energy. **(c)** Explain why this film can be handled in a darkroom under red light.

[6.83] In an experiment to study the photoelectric effect, a scientist measures the kinetic energy of ejected electrons as a function of the frequency of radiation hitting a metal surface. She obtains the following plot:



The point labeled "ν_0" corresponds to light with a wavelength of 680 nm. **(a)** What is the value of ν_0 in s^{-1}? **(b)** What is the value of the work function of the metal in units of kJ/mol of ejected electrons? **(c)** What happens when the metal is irradiated with light of frequency less than ν_0? **(d)** Note that when the frequency of the light is greater than ν_0, the plot shows a straight line with a nonzero slope. Why is this the case? **(e)** Can you determine the slope of the line segment discussed in part (d)? Explain.

6.84 The series of emission lines of the hydrogen atom for which $n_f = 3$ is called the *Paschen series*. **(a)** Determine the region of the electromagnetic spectrum in which the lines of the Paschen series are observed. **(b)** Calculate the wavelengths of the first three lines in the Paschen series—those for which $n_i = 4, 5$, and 6.

6.85 **CQ** When the spectrum of light from the Sun is examined in high resolution in an experiment similar to that illustrated in Figure 6.10, dark lines are evident. These are called Fraunhofer lines, after the scientist who studied them extensively in the early nineteenth century. Altogether, about 25,000 lines have been identified in the solar spectrum between 2950 Å and 10,000 Å. The Fraunhofer lines are attributed to absorption of certain wavelengths of the Sun's "white" light by gaseous elements in the Sun's atmosphere. **(a)** Describe the process that causes absorption of specific wavelengths of light from the solar spectrum. **(b)** If a scientist wanted to know which Fraunhofer lines belonged to a given element, say neon, what experiments could she conduct here on Earth to provide data?

[6.86] Bohr's model can be used for hydrogen-like ions—ions that have only one electron, such as He^+ and Li^{2+}. **(a)** Why is the Bohr model applicable to He^+ ions but not to neutral He atoms? **(b)** The ground-state energies of H, He^+, and Li^{2+} are tabulated as follows:

Atom or ion	H	He^+	Li^{2+}
Ground-state energy	-2.18×10^{-18} J	-8.72×10^{-18} J	-1.96×10^{-17} J

By examining these numbers, propose a relationship between the ground-state energy of hydrogen-like systems and the nuclear charge, *Z*. **(c)** Use the relationship you derive in part (b) to predict the ground-state energy of the C^{5+} ion.

6.87 Under appropriate conditions, molybdenum emits X rays that have a characteristic wavelength of 0.711 Å. These X rays are used in diffraction experiments to determine the structures of molecules. How fast would an electron have to be moving in order to have the same wavelength as these X rays?

[6.88] An electron is accelerated through an electric potential to a kinetic energy of 18.6 keV. What is its characteristic wavelength? [*Hint:* Recall that the kinetic energy of a moving object is $E = \frac{1}{2}mv^2$, where *m* is the mass of the object and *v* is the speed of the object.]

6.89 What is the difference between an *orbit* (Bohr model of the hydrogen atom) and an *orbital* (quantum mechanical model of the hydrogen atom)?

6.90 **CQ** Which of the quantum numbers governs **(a)** the shape of an orbital, **(b)** the energy of an orbital, **(c)** the spin properties of the electron, **(d)** the spatial orientation of the orbital?

[6.91] Consider the discussion of radial probability functions in the "A Closer Look" box in Section 6.6. **(a)** What is the difference between the probability density as a function of *r* and the radial probability function as a function of *r*? **(b)** What is the significance of the term $4\pi r^2$ in the radial probability functions for the *s* orbitals? **(c)** Based on Figures 6.18 and 6.21, make sketches of what you think the probability density as a function of *r* and the radial probability function would look like for the 4*s* orbital of the hydrogen atom.

6.92 The "magic numbers" in the periodic table are the atomic numbers of elements with high stability (the noble gases): 2, 10, 18, 36, 54, and 86. In terms of allowed values of orbitals and spin quantum numbers, explain why these electron arrangements correspond to special stability.

[6.93] For non-spherically symmetric orbitals, the contour representations (as in Figures 6.22 and 6.23) suggest where nodal planes exist (that is, where the electron density is zero). For example, the p_x orbital has a node wherever $x = 0$; this equation is satisfied by all points on the *yz* plane, so this plane is called a nodal plane of the p_x orbital. **(a)** Determine the nodal plane of the p_z orbital. **(b)** What are the two nodal planes of the d_{xy} orbital? **(c)** What are the two nodal planes of the $d_{x^2-y^2}$ orbital?

[6.94] As noted in Figure 6.25, the spin of an electron generates a magnetic field, with spin-up and spin-down electrons having opposite fields. In the absence of a magnetic field, a spin-up and a spin-down electron have the same energy. **(a)** Why do you think that the use of a magnet was important in the discovery of electron spin (see the "A Closer Look" box in Section 6.7)? **(b)** Imagine that the two spinning electrons in Figure 6.25 were placed between the poles of a horseshoe magnet, with the north pole of the magnet at the top of the figure. Based on what you know about magnets, would you expect the left or right electron in the figure to have the lower energy? **(c)** A phenomenon called *electron spin resonance* (ESR) is closely related to nuclear magnetic resonance. In ESR a compound with an unpaired electron is placed in a magnetic field, which causes the unpaired electron to have two different energy states analogous to Figure 6.27. ESR uses microwave radiation to excite the unpaired electron from one state to the other. Based on your reading of the "Chemistry and Life" box in Section 6.7, does an ESR experiment require photons of greater or lesser energy than an NMR experiment?

[6.95] The "Chemistry and Life" box in Section 6.7 described the techniques called NMR and MRI. **(a)** Instruments for obtaining MRI data are typically labeled with a frequency, such as 600 MHz. Why do you suppose this label is relevant to the experiment? **(b)** What is the value of ΔE in Figure 6.27 that would correspond to the absorption of a photon of radiation with frequency 450 MHz? **(c)** In general, the stronger the magnetic field, the greater the information obtained from an NMR or MRI experiment. Why do you suppose this is the case?

[6.96] **CQ** Suppose that the spin quantum number, m_s, could have *three* allowed values instead of two. How would this affect the number of elements in the first four rows of the periodic table?

6.97 Using only a periodic table as a guide, write the condensed electron configurations for the following atoms: **(a)** Se, **(b)** Rh, **(c)** Si, **(d)** Hg, **(e)** Hf.

6.98 Scientists have speculated that element 126 might have a moderate stability, allowing it to be synthesized and characterized. Predict what the condensed electron configuration of this element might be.

Integrative Exercises

[6.99] Microwave ovens use microwave radiation to heat food. The microwaves are absorbed by moisture in the food, which is transferred to other components of the food. As the water becomes hotter, so does the food. Suppose that the microwave radiation has a wavelength of 11.2 cm. How many photons are required to heat 200 mL of coffee from 23°C to 60°C?

6.100 The stratospheric ozone (O_3) layer helps to protect us from harmful ultraviolet radiation. It does so by absorbing ultraviolet light and falling apart into an O_2 molecule and an oxygen atom, a process known as photodissociation.

$$O_3(g) \longrightarrow O_2(g) + O(g)$$

Use the data in Appendix C to calculate the enthalpy change for this reaction. What is the maximum wavelength a photon can have if it is to possess sufficient energy to cause this dissociation? In what portion of the spectrum does this wavelength occur?

6.101 The discovery of hafnium, element number 72, provided a controversial episode in chemistry. G. Urbain, a French chemist, claimed in 1911 to have isolated an element number 72 from a sample of rare earth (elements 58–71) compounds. However, Niels Bohr believed that hafnium was more likely to be found along with zirconium than with the rare earths. D. Coster and G. von Hevesy, working in Bohr's laboratory in Copenhagen, showed in 1922 that element 72 was present in a sample of Norwegian zircon, an ore of zirconium. (The name hafnium comes from the Latin name for Copenhagen, *Hafnia*). **(a)** How would you use electron configuration arguments to justify Bohr's prediction? **(b)** Zirconium, hafnium's neighbor in group 4B, can be produced as a metal by reduction of solid $ZrCl_4$ with molten sodium metal. Write a balanced chemical equation for the reaction. Is this an oxidation-reduction reaction? If yes, what is reduced and what is oxidized? **(c)** Solid zirconium dioxide, ZrO_2, is reacted with chlorine gas in the presence of carbon. The products of the reaction are $ZrCl_4$ and two gases, CO_2 and CO in the ratio 1:2. Write a balanced chemical equation for the reaction. Starting with a 55.4-g sample of ZrO_2, calculate the mass of $ZrCl_4$ formed, assuming that ZrO_2, is the limiting reagent and assuming 100% yield. **(d)** Using their electron configurations, account for the fact that Zr and Hf form chlorides MCl_4 and oxides MO_2.

6.102 **(a)** Account for formation of the following series of oxides in terms of the electron configurations of the elements and the discussion of ionic compounds in Section 2.7: $K_2O, CaO, Sc_2O_3, TiO_2, V_2O_5, CrO_3$. **(b)** Name these oxides. **(c)** Consider the metal oxides whose enthalpies of formation (in kJ mol^{-1}) are listed here.

Oxide	$K_2O(s)$	$CaO(s)$	$TiO_2(s)$	$V_2O_5(s)$
ΔH_f°	−363.2	−635.1	−938.7	−1550.6

Calculate the enthalpy changes in the following general reaction for each case:

$$M_nO_m(s) + H_2(g) \longrightarrow nM(s) + mH_2O(g)$$

(You will need to write the balanced equation for each case, then compute $\Delta H°$.) **(d)** Based on the data given, estimate a value of ΔH_f° for $Sc_2O_3(s)$.

6.103 CQ The first 25 years of the twentieth century were momentous for the rapid pace of change in scientists' understanding of the nature of matter. **(a)** How did Rutherford's experiments on the scattering of α particles by a gold foil set the stage for Bohr's theory of the hydrogen atom? **(b)** In what ways is de Broglie's hypothesis, as it applies to electrons, consistent with J. J. Thomson's conclusion that the electron has mass? In what sense is it consistent with proposals that preceded Thomson's work, that the cathode rays are a wave phenomenon?

[6.104] The two most common isotopes of uranium are ^{235}U and ^{238}U. **(a)** Compare the number of protons, the number of electrons, and the number of neutrons in atoms of these two isotopes. **(b)** Using the periodic table in the front inside cover, write the electron configuration for a U atom. **(c)** Compare your answer to part (b) to the electron configuration given in Figure 6.30. How can you explain any differences between these two electron configurations? **(d)** ^{238}U undergoes radioactive decay to ^{234}Th. How many protons, electrons, and neutrons are gained or lost by the ^{238}U atom during this process? **(e)** Examine the electron configuration for Th in Figure 6.30. Are you surprised by what you find? Explain.

eMEDIA EXERCISES

These exercises make use of the interactive objects available online in OneKey or the Companion Website, and on your Accelerator CD. Access to these resources comes in your MediaPak.

6.105 The **Electromagnetic Spectrum** activity (*6.1*) allows you to choose a color in the visible spectrum and see its wavelength, frequency, and energy per photon. **(a)** What is the wavelength range of blue light? **(b)** What are the ranges of its frequency and energy per photon? **(c)** Exercise 6.25 indicates that a type of sunburn is caused by light with wavelength ~325 nm. Would you expect any of the visible wavelengths to cause sunburn? Explain.

6.106 In the **Flame Tests for Metals** movie (*6.3*) the characteristic color of the flame is produced by emissions at several visible wavelengths, with the most intense spectral lines dominating the color. For instance, the most intense visible lines in the spectrum of lithium occur at ~671 nm. **(a)** What color is light of this wavelength? **(b)** At what approximate wavelength would you expect to find the most intense lines in the visible spectrum of potassium? **(c)** Based on the movie, how would you expect the intensity of visible lines in the spectrum of potassium to compare to those in the spectrum of lithium? **(d)** Would it be possible to verify the