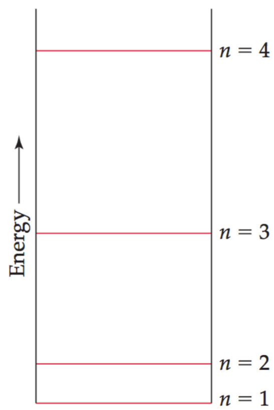
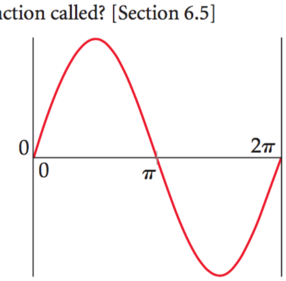
Electronic Structure of Atoms Problem Set

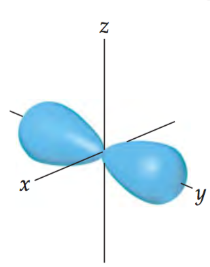
6.6 A certain quantum mechanical system has the energy levels shown in the diagram below. The energy levels are indexed by a single quantum number n that is an integer. **(a)** As drawn, which quantum numbers are involved in the transition that requires the most energy? **(b)** Which quantum numbers 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]



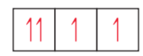
6.7 Consider a fictitious one-dimensional system with one electron. The wave function for the electron, drawn at the top of the next page, is $\psi(x) = \sin x$ from $x = 0$ to $x = 2\pi$. **(a)** Sketch the probability density, $\psi^2(x)$, from $x = 0$ to $x = 2\pi$. **(b)** At what value or values of x will there be the greatest probability of finding the electron? **(c)** What is the probability that the electron will be found at $x = \pi$? What is such a point in a wave function called? [Section 6.5]



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



6.9 The drawing below shows part of the orbital diagram for an element. **(a)** As drawn, the drawing is *incorrect*. Why? **(b)** How would you correct the drawing without changing the number of electrons? **(c)** To which group in the periodic table does the element belong? [Section 6.8]



6.10 State where in the periodic table these elements appear:

(a) elements with the valence-shell electron configuration ns^2np^5

(b) elements that have three unpaired p electrons

(c) an element whose valence electrons are $4s^24p^1$

(d) the d-block elements

6.13 Label each of the following statements as true or false. For those that are false, correct the statement. **(a)** Visible light is a form of electromagnetic radiation. **(b)** Ultraviolet light has longer wavelengths than visible light. **(c)** X-rays travel faster than microwaves. **(d)** Electromagnetic radiation and sound waves travel at the same speed.

6.14 Determine which of the following statements are false and correct them. **(a)** The frequency of radiation increases as the wavelength increases. **(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.

6.17 **(a)** What is the frequency of radiation that has a wavelength of 10 μm, about the size of a bacterium? **(b)** What is the wavelength of radiation that has a frequency of 5.50×10^{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?

6.23 **(a)** Calculate the energy of a photon of electromagnetic radiation whose frequency is 6.75×10^{12} s^{-1}. (b) Calculate the energy of a photon of radiation whose wavelength is 322 nm. (c) What wavelength of radiation has photons of energy 2.87×10^{-18} J?

6.29 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?

6.32 Sodium metal requires a photon with a minimum energy of 4.41×10^{-19} J to emit electrons. **(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 405 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 ?

6.34 **(a)** In terms of the Bohr theory of the hydrogen atom, what 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?

6.35 Is energy emitted or absorbed when the following electronic 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?

6.38 **(a)** Calculate the energies of an electron in the hydrogen atom for $n = 1$ and for $n = \infty$. How much energy does it require to move the electron out of the atom completely (from $n = 1$ to $n = \infty$), according to Bohr? Put your answer in kJ/mol. **(b)** The energy for the process $H + \text{energy} \rightarrow H^+ + e^-$ is called the ionization energy of hydrogen. The experimentally determined value for the ionization energy of hydrogen is 1310 kJ/mol. How does this compare to your calculation?

6.41 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.

6.43 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, **(d)** an ozone (O_3) molecule in the upper atmosphere moving at 550 m/s.

6.46 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 8.95×10^6 m/s. What is the characteristic wavelength of this electron? Is the wavelength comparable to the size of atoms?

6.48 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?

6.49 **(a)** Why does the Bohr model of the hydrogen atom violate 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?

6.51 **(a)** For $n = 4$, what are the possible values of l? **(b)** For $l = 2$, what are the possible values of m_l? **(c)** If m_l is 2, what are the possible values for l?

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

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

6.61 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?

6.62 **(a)** The average distance from the nucleus of a 3*s* electron in a chlorine atom is smaller than that for a 3*p* 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 3*s* electron from the chlorine atom, as compared with a 2*p* electron? Explain.

6.66 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)** $n = 4$, $l = 0$, $m_l = 0$?

6.67 **(a)** What are "valence electrons"? **(b)** What are "core electrons"? **(c)** What does each box in an orbital diagram represent? **(d)** What quantity is represented by the half arrows in an orbital diagram?

6.68 For each element, indicate the number of valence electrons, core electrons, and unpaired electrons in the ground state: **(a)** carbon, **(b)** phosphorus, **(c)** neon.

6.69 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)** U, **(f)** Pb.

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

6.73 What is wrong with the following electron configurations for atoms in their ground states? **(a)** $ls^2 2s^2 3s^1$, **(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)** $[Kr]4d^6 5s^2 5p^1$.