Solutions to suggested end-of-chapter problems for Ch. 7

Question numbers correspond to the 9^{th} edition. The number in parentheses corresponds to the 10^{th} edition.

45 (51). The wavelength is the distance between consecutive wave peaks. Wave *a* shows 4 wave-lengths, and wave *b* shows 8 wavelengths.

Wave *a*:
$$\lambda = \frac{1.6 \times 10^{-3} \text{ m}}{4} = 4.0 \times 10^{-4} \text{ m}$$

Wave *b*: $\lambda = \frac{1.6 \times 10^{-3} \text{ m}}{8} = 2.0 \times 10^{-4} \text{ m}$

Wave a has the longer wavelength. Because frequency and photon energy are both inversely proportional to wavelength, wave b will have the higher frequency and larger photon energy since it has the shorter wavelength.

$$\nu = \frac{c}{\lambda} = \frac{2.998 \times 10^8 \text{ m/s}}{2.0 \times 10^{-4} \text{ m}} = 1.5 \times 10^{12} \text{ s}^{-1}$$
$$E = \frac{hc}{\lambda} = \frac{6.626 \times 10^{-34} \text{ J} \text{ s} \times 2.998 \times 10^8 \text{ m/s}}{2.0 \times 10^{-4} \text{ m}} = 9.9 \times 10^{-22} \text{ J}$$

Because both waves are examples of electromagnetic radiation, both waves travel at the same speed, c, the speed of light. From Figure 7.2 of the text, both of these waves represent infrared electromagnetic radiation.

46 (52). Referencing Figure 7.2 of the text, 2.12×10^{-10} m electromagnetic radiation is X rays.

$$\lambda = \frac{c}{v} = \frac{2.9979 \times 10^8 \text{ m/s}}{107.1 \times 10^6 \text{ s}^{-1}} = 2.799 \text{ m}$$

From the wavelength calculated above, 107.1 MHz electromagnetic radiation is FM radio-waves.

$$\lambda = \frac{hc}{E} = \frac{6.626 \times 10^{-34} \text{ J s} \times 2.998 \times 10^8 \text{ m/s}}{3.97 \times 10^{-19} \text{ J}} = 5.00 \times 10^{-7} \text{ m}$$

The 3.97×10^{-19} J/photon electromagnetic radiation is visible (green) light.

The photon energy and frequency order will be the exact opposite of the wavelength ordering because E and v are both inversely related to λ . From the previously calculated wavelengths, the order of photon energy and frequency is:

FM radiowaves	<	visible (green) light	<	X rays
longest λ				shortest $\boldsymbol{\lambda}$
lowest v				highest v
smallest E				largest E

47 (53).
$$E_{\text{photon}} = \frac{hc}{\lambda} = \frac{6.626 \times 10^{-34} \text{ J s} \times 2.998 \times 10^8 \text{ m/s}}{150. \text{ nm} \times \frac{1 \text{ m}}{1 \times 10^9 \text{ nm}}} = 1.32 \times 10^{-18} \text{ J}$$

$$1.98 \times 10^5 \text{ J} \times \frac{1 \text{ photon}}{1.32 \times 10^{-18} \text{ J}} \times \frac{1 \text{ atom C}}{\text{ photon}} = 1.50 \times 10^{23} \text{ atoms C}$$

50 (56).
$$\frac{208.4 \text{ kJ}}{\text{mol}} \times \frac{1 \text{ mol}}{6.0221 \times 10^{23}} = 3.461 \times 10^{-22} \text{ kJ} = 3.461 \times 10^{-19} \text{ J to remove one electron}$$

$$E = \frac{hc}{\lambda}, \ \lambda = \frac{hc}{E} = \frac{6.6261 \times 10^{-34} \text{ J s} \times 2.9979 \times 10^8 \text{ m/s}}{3.461 \times 10^{-19} \text{ J}} = 5.739 \times 10^{-7} \text{ m} = 573.9 \text{ nm}$$

51 (57). Ionization energy = energy to remove an electron = $7.21 \times 10^{-19} = E_{\text{photon}}$

$$E_{photon} = hv \text{ and } \lambda v = c. \text{ So } v = \frac{c}{\lambda} \text{ and } E = \frac{hc}{\lambda}.$$
$$\lambda = \frac{hc}{E_{photon}} = \frac{6.626 \times 10^{-34} \text{ J s} \times 2.998 \times 10^8 \text{ m/s}}{7.21 \times 10^{-19} \text{ J}} = 2.76 \times 10^{-7} \text{ m} = 276 \text{ nm}$$

52 (58).
$$\frac{890.1 \text{ kJ}}{\text{mol}} \times \frac{1 \text{ mol}}{6.0221 \times 10^{23} \text{ atoms}} = \frac{1.478 \times 10^{-21} \text{ kJ}}{\text{atom}} = \frac{1.478 \times 10^{-18} \text{ J}}{\text{atom}}$$

= ionization energy per atom

$$E = \frac{hc}{\lambda}, \ \lambda = \frac{hc}{E} = \frac{6.626 \times 10^{-34} \text{ J s} \times 2.9979 \times 10^8 \text{ m/s}}{1.478 \times 10^{-18} \text{ J}} = 1.344 \times 10^{-7} \text{ m} = 134.4 \text{ nm}$$

No, it will take light having a wavelength of 134.4 nm or less to ionize gold. A photon of light having a wavelength of 225 nm is longer wavelength and thus lower energy than 134.4 nm light.

53 (59). a. 10.% of speed of light = $0.10 \times 3.00 \times 10^8$ m/s = 3.0×10^7 m/s

$$\lambda = \frac{h}{mv}, \ \lambda = \frac{6.63 \times 10^{-34} \text{ J s}}{9.11 \times 10^{-31} \text{ kg } \times 3.0 \times 10^7 \text{ m/s}} = 2.4 \times 10^{-11} \text{ m} = 2.4 \times 10^{-2} \text{ nm}$$

Note: For units to come out, the mass must be in kg because $1 \text{ J} = \frac{1 \text{ kg m}^2}{\text{s}^2}$.

b.
$$\lambda = \frac{h}{mv} = \frac{6.63 \times 10^{-34} \text{ J s}}{0.055 \text{ kg} \times 35 \text{ m/s}} = 3.4 \times 10^{-34} \text{ m} = 3.4 \times 10^{-25} \text{ nm}$$

This number is so small that it is insignificant. We cannot detect a wavelength this small. The meaning of this number is that we do not have to worry about the wave properties of large objects.

64 (70). a. False; it takes less energy to ionize an electron from n = 3 than from the ground state.

- b. True
- c. False; the energy difference between n = 3 and n = 2 is smaller than the energy difference between n = 3 and n = 1; thus the wavelength is larger for the $n = 3 \rightarrow n = 2$ electronic transition than for the $n = 3 \rightarrow n = 1$ transition. E and λ are inversely proportional to each other (E = hc/ λ).
- d. True
- e. False; n = 2 is the first excited state, and n = 3 is the second excited state.

69 (75). a.
$$\Delta p = m\Delta v = 9.11 \times 10^{-31} \text{ kg} \times 0.100 \text{ m/s} = \frac{9.11 \times 10^{-32} \text{ kg m}}{\text{s}}$$

 $\Delta p\Delta x \ge \frac{h}{4\pi}, \ \Delta x = \frac{h}{4\pi\Delta p} = \frac{6.626 \times 10^{-34} \text{ J s}}{4 \times 3.142 \times (9.11 \times 10^{-32} \text{ kg m/s})} = 5.79 \times 10^{-4} \text{ m}$
b. $\Delta x = \frac{h}{4\pi\Delta p} = \frac{6.626 \times 10^{-34} \text{ J s}}{4 \times 3.142 \times 0.145 \text{ kg} \times 0.100 \text{ m/s}} = 3.64 \times 10^{-33} \text{ m}$

- c. The diameter of an H atom is roughly $\sim 10^{-8}$ cm. The uncertainty in position is much larger than the size of the atom.
- d. The uncertainty is insignificant compared to the size of a baseball.
- 70 (76). Units of $\Delta E \cdot \Delta t = J \times s$, the same as the units of Planck's constant.

Units of
$$\Delta(mv) \cdot \Delta x = kg \times \frac{m}{s} \times m = \frac{kg m^2}{s} = \frac{kg m^2}{s^2} \times s = J \times s$$

73 (79). a. allowed b. For $\ell = 3$, m_{ℓ} can range from -3 to +3; thus +4 is not allowed.

c. *n* cannot equal zero. d. ℓ cannot be a negative number.

74 (80). a. For n = 3, $\ell = 3$ is not possible.

- d. m_s cannot equal -1.
- e. ℓ cannot be a negative number.
- f. For $\ell = 1$, m_{ℓ} cannot equal 2.

The quantum numbers in parts b and c are allowed.

84 (90). The two exceptions are Cr and Cu.

Cr: $1s^22s^22p^63s^23p^64s^13p^5$; Cr has 6 unpaired electrons.





85 (91). Si: 1s²2s²2p⁶3s²3p² or [Ne]3s²3p²; Ga: 1s²2s²2p⁶3s²3p⁶4s²3d¹⁰4p¹ or [Ar]4s²3d¹⁰4p¹
As: [Ar]4s²3d¹⁰4p³; Ge: [Ar]4s²3d¹⁰4p²; Al: [Ne]3s²3p¹; Cd: [Kr]5s²4d¹⁰
S: [Ne]3s²3p⁴; Se: [Ar]4s²3d¹⁰4p⁴

90 (96). a. As:
$$1s^22s^22p^63s^23p^64s^23d^{10}4p^3$$

- b. Element 116 will be below Po in the periodic table: $[Rn]7s^25f^{14}6d^{10}7p^4$
- c. Ta: $[Xe]6s^{2}4f^{14}5d^{3}$ or Ir: $[Xe]6s^{2}4f^{14}5d^{7}$
- d. At: $[Xe]6s^24f^{14}5d^{10}6p^5$; note that element 117 will also have electrons in the 6p atomic orbitals (as well as electrons in the 7p orbitals).
- **91 (97).** a. The complete ground state electron for this neutral atom is $1s^22s^22p^63s^23p^4$. This atom has 2 + 2 + 6 + 2 + 4 = 16 electrons. Because the atom is neutral, it also has 16 protons, making the atom sulfur, S.
 - b. Complete excited state electron configuration: $1s^22s^12p^4$; this neutral atom has 2 + 1 + 4 = 7 electrons, which means it has 7 protons, which identifies it as nitrogen, N.
 - c. Complete ground state electron configuration: 1s²2s²2p⁶3s²3p⁶4s²3d¹⁰4p⁵; this 1- charged ion has 35 electrons. Because the overall charge is 1-, this ion has 34 protons which identifies it as selenium. The ion is Se⁻.
- **92 (98).** a. This atom has 10 electrons. Ne b. S
 - c. The predicted ground state configuration is $[Kr]5s^24d^9$. From the periodic table, the element is Ag. *Note*: $[Kr]5s^14d^{10}$ is the actual ground state electron configuration for Ag.

d. Bi: $[Xe]6s^24f^{14}5d^{10}6p^3$; the three unpaired electrons are in the 6p orbitals.

93 (99). Hg: $1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^24d^{10}5p^66s^24f^{14}5d^{10}$

- a. From the electron configuration for Hg, we have $3s^2$, $3p^6$, and $3d^{10}$ electrons; 18 total electrons with n = 3.
- b. $3d^{10}$, $4d^{10}$, $5d^{10}$; 30 electrons are in the d atomic orbitals.
- c. $2p^6$, $3p^6$, $4p^6$, $5p^6$; each set of *n*p orbitals contain one p_z atomic orbital. Because we have 4 sets of *n*p orbitals and two electrons can occupy the p_z orbital, there are 4(2) = 8 electrons in p_z atomic orbitals.
- d. All the electrons are paired in Hg, so one-half of the electrons are spin up $(m_s = +1/2)$ and the other half are spin down $(m_s = -1/2)$. 40 electrons have spin up.
- 100 (106). The number of unpaired electrons is in parentheses.

a.	ed state of boron	(1)	b.	ground state of neon	(0)
	ound state: 1s ² 2s ² 2p ¹	(1)		Ne ground state: $1s^22s^22p^6$	(0)
c.	d state of fluorine	(3)	d.	excited state of iron	(6)
	ound state: 1s ² 2s ² 2p ⁵	(1)		Fe ground state: [Ar]4s ² 3d ⁶	(4)
	$\downarrow \uparrow \downarrow \uparrow$			$\stackrel{\uparrow\downarrow}{\downarrow} \stackrel{\uparrow}{=} \stackrel{\downarrow}{=} \stackrel{\downarrow}{=} \stackrel{\downarrow}{=} \stackrel{\uparrow}{=} \stackrel{\uparrow}{=} \stackrel{\uparrow}{=} \stackrel{\downarrow}{=} $	
	$\downarrow \uparrow \downarrow \uparrow _{2p}$			$ \begin{array}{c} \uparrow \downarrow \uparrow _ \uparrow _ \uparrow _ \uparrow \\ 3d \end{array} \begin{array}{c} \uparrow _ \uparrow \\ 3d \end{array}$	

104 (110). We get the number of unpaired electrons by examining the incompletely filled subshells.

O: $[He]2s^22p^4$	2p ⁴ :	$\uparrow \downarrow \uparrow _ \uparrow _$	two unpaired e
$O^+: [He] 2s^2 2p^3$	2p ³ :		three unpaired e ⁻
$O^{-}: [He] 2s^2 2p^5$	2p ⁵ :		one unpaired e ⁻
Os: $[Xe]6s^24f^{14}5d^6$	5d ⁶ :	$\uparrow \downarrow \uparrow \downarrow \uparrow \uparrow \uparrow \uparrow$	four unpaired e ⁻
$Zr: [Kr]5s^24d^2$	$4d^2$:	î_ î	two unpaired e ⁻
S: $[Ne]3s^23p^4$	3p ⁴ :	$\uparrow \downarrow \uparrow _ \uparrow _$	two unpaired e ⁻
F: $[He]2s^22p^5$	2p ⁵ :		one unpaired e ⁻
Ar: [Ne]3s ² 3p ⁶	3p ⁶	$\uparrow \downarrow \uparrow \downarrow \uparrow \downarrow$	zero unpaired e ⁻

105 (111). Size (radius) decreases left to right across the periodic table, and size increases from top to bottom of the periodic table.

a. S < Se < Te b. Br < Ni < K c. F < Si < Ba

All follow the general radius trend.

106 (112). a. Be < Na < Rb b. Ne < Se < Sr c. O < P < Fe

All follow the general radius trend.

- **107** (**113**). The ionization energy trend is the opposite of the radius trend; ionization energy (IE), in general, increases left to right across the periodic table and decreases from top to bottom of the periodic table.
 - a. Te < Se < S b. K < Ni < Br c. Ba < Si < F

All follow the general ionization energy trend.

108 (114). a. Rb < Na < Be b. Sr < Se < Ne c. Fe < P < O

All follow the general ionization energy trend.

- **109** (115). a. He (From the general radius trend.) b. Cl
 - c. Element 116 is the next oxygen family member to be discovered (under Po), element 119 is the next alkali metal to be discovered (under Fr), and element 120 is the next alkaline earth metal to be discovered (under Ra). From the general radius trend, element 116 will be the smallest.
 - d. Si
 - e. Na⁺; this ion has the fewest electrons as compared to the other sodium species present. Na⁺ has the smallest number of electron-electron repulsions, which makes it the smallest ion with the largest ionization energy.
- **110** (116). a. Ba (From the general ionization energy trend.) b. K
 - c. O; in general, Group 6A elements have a lower ionization energy than neighboring Group 5A elements. This is an exception to the general ionization energy trend across the periodic table.
 - d. S^{2-} ; this ion has the most electrons compared to the other sulfur species present. S^{2-} has the largest number of electron-electron repulsions, which leads to S^{2-} having the largest size and smallest ionization energy.
 - e. Cs; this follows the general ionization energy trend.
- **115 (121).** a. As we remove succeeding electrons, the electron being removed is closer to the nucleus, and there are fewer electrons left repelling it. The remaining electrons are more strongly attracted to the nucleus, and it takes more energy to remove these electrons.

b. Al: $1s^22s^22p^63s^23p^1$; for I₄, we begin removing an electron with n = 2. For I₃, we remove an electron with n = 3 (the last valence electron). In going from n = 3 to n = 2, there is a big jump in ionization energy because the n = 2 electrons are closer to the nucleus on average than the n = 3 electrons. Since the n = 2 electrons are closer, on average, to the nucleus, they are held more

tightly and require a much larger amount of energy to remove compared to the n = 3 electrons. In general, valence electrons are much easier to remove than inner-core electrons.

- **118 (124).** a. More favorable electron affinity: K and Cl; Mg has a positive electron affinity value, and F has a more positive electron affinity value than expected from its position relative to Cl.
 - b. Higher ionization energy: Mg and F c. Larger radius: K and Cl
- **121 (127).** The electron affinity trend is very erratic. In general, electron affinity decreases down the periodic table, and the trend across the table is too erratic to be of much use.

127 (135). It should be potassium peroxide (K_2O_2) because K⁺ ions are stable in ionic compounds. K²⁺ ions are not stable; the second ionization energy of K is very large compared to the first.

133 (143). a.
$$6 \operatorname{Li}(s) + N_2(g) \rightarrow 2 \operatorname{Li}_3 N(s)$$
 b. $2 \operatorname{Rb}(s) + S(s) \rightarrow \operatorname{Rb}_2 S(s)$

138 (148). Energy to make water boil = $s \times m \times \Delta T = \frac{4.18 \text{ J}}{^{\circ}\text{C g}} \times 50.0 \text{ g} \times 75.0^{\circ}\text{C} = 1.57 \times 10^{4} \text{ J}$

$$E_{\text{photon}} = \frac{\text{hc}}{\lambda} = \frac{6.626 \times 10^{-34} \text{ J s} \times 2.998 \times 10^8 \text{ m/s}}{9.75 \times 10^{-2} \text{ m}} = 2.04 \times 10^{-24} \text{ J}$$

$$1.57 \times 10^4 \text{ J} \times \frac{1 \text{ s}}{750 \text{ J}} = 20.9 \text{ s}; \ 1.57 \times 10^4 \text{ J} \times \frac{1 \text{ photon}}{2.04 \times 10^{-24} \text{ J}} = 7.70 \times 10^{27} \text{ photons}$$

143 (153). a. True for H only. b. True for all atoms. c. True for all atoms.

- **145 (155).** 1p: n = 1, $\ell = 1$ is not possible; 3f: n = 3, $\ell = 3$ is not possible; 2d: n = 2, $\ell = 2$ is not possible; in all three incorrect cases, $n = \ell$. The maximum value ℓ can have is n 1, not n.
- **148 (158).** a. The 4+ ion contains 20 electrons. Thus the electrically neutral atom will contain 24 electrons. The atomic number is 24, which identifies it as chromium.
 - b. The ground state electron configuration of the ion must be $1s^22s^22p^63s^23p^64s^03d^2$; there are 6 electrons in s orbitals.
 - c. 12
 - d. 2
 - e. From the mass, this is the isotope ${}^{50}_{24}$ Cr. There are 26 neutrons in the nucleus.
 - f. $1s^22s^22p^63s^23p^64s^13d^5$ is the ground state electron configuration for Cr. Cr is an exception to the normal filling order.

149 (159). Valence electrons are easier to remove than inner-core electrons. The large difference in energy between I_2 and I_3 indicates that this element has two valence electrons. This element is most likely an alkaline earth metal since alkaline earth metal elements all have two valence electrons.