

Chapter 6: Electronic Structure and the Periodic Table

2. Refer to Section 6.1 and Examples 6.1 and 6.2.

(a) $c = \lambda\nu$

$$2.998 \times 10^8 \text{ m/s} = \lambda(4.00 \times 10^8/\text{s})$$

$$\lambda = 0.750 \text{ m}$$

(b) $E = h\nu = (6.626 \times 10^{-34} \text{ J}\cdot\text{s})(4.00 \times 10^8/\text{s}) = 2.65 \times 10^{-25} \text{ J}$

Note that this is for one photon, thus units are J/photon.

(c) $\frac{2.65 \times 10^{-25} \text{ J}}{1 \text{ photon}} \times \frac{6.02 \times 10^{23} \text{ photons}}{1 \text{ mol.}} = 1.60 \times 10^{-1} \text{ J/mol.} = 1.60 \times 10^{-4} \text{ kJ/mol.}$

4. Refer to Section 6.1, Example 6.1, and Figure 6.2.

(a) $1498 \text{ nm} \times \frac{1 \text{ m}}{1 \times 10^9 \text{ nm}} = 1.498 \times 10^{-6} \text{ m}$

$1.498 \times 10^{-6} \text{ m}$ is just outside the visible region in the **infrared** region.

(b) $c = \lambda\nu$

$$\nu = \frac{2.998 \times 10^8 \text{ m/s}}{1498 \times 10^{-6} \text{ m}} = 2.001 \times 10^{14} \text{ s}^{-1} = 2.001 \times 10^{14} \text{ Hz}$$

(c) $E = h\nu = (6.626 \times 10^{-34} \text{ J}\cdot\text{s})(2.001 \times 10^{14} \text{ s}^{-1}) = 1.326 \times 10^{-19} \text{ J}$

6. Refer to Example 6.1.

$$\frac{941 \text{ kJ}}{1 \text{ mol.}} \times \frac{1000 \text{ J}}{1 \text{ kJ}} \times \frac{1 \text{ mol.}}{6.022 \times 10^{23} \text{ photons}} = 1.56 \times 10^{-18} \text{ J}$$

$$\lambda = \frac{hc}{E} = \frac{(6.626 \times 10^{-34} \text{ J}\cdot\text{s})(2.998 \times 10^8 \text{ m/s})}{1.56 \times 10^{-18} \text{ J}} = 1.27 \times 10^{-7} \text{ m} = 127 \text{ nm}$$

8. Refer to Section 6.1 and Example 6.2.

Calculate the amount of energy in one photon. Use that and the total amount of energy to calculate the number of photons emitted.

$$\lambda = 633 \text{ nm} \times \frac{1 \text{ m}}{1 \times 10^9 \text{ nm}} = 6.33 \times 10^{-7} \text{ m}$$

$$E = \frac{hc}{\lambda} = \frac{(6.626 \times 10^{-34} \text{ J} \cdot \text{s})(2.998 \times 10^8 \text{ m/s})}{6.33 \times 10^{-7} \text{ m}} = 3.14 \times 10^{-19} \text{ J}$$

$$12 \text{ kJ} \times \frac{1000 \text{ J}}{1 \text{ kJ}} \times \frac{1 \text{ photon}}{3.14 \times 10^{-19} \text{ J}} = 3.8 \times 10^{22} \text{ photons}$$

10. Refer to Section 6.2, Example 6.3, and Figure 6.2.

Use the Rydberg equation to calculate the frequency. Then calculate the wavelength to determine the region of the spectrum associated with that frequency.

$$\nu = \frac{R_H}{h} \left[\frac{1}{(n_{lo})^2} - \frac{1}{(n_{hi})^2} \right] = \frac{2.180 \times 10^{-18} \text{ J}}{6.626 \times 10^{-34} \text{ J} \cdot \text{s}} \left[\frac{1}{1^2} - \frac{1}{3^2} \right] = 2.925 \times 10^{15} \text{ s}^{-1}$$

$$\lambda = \frac{c}{\nu} = \frac{2.998 \times 10^8 \text{ m/s}}{2.925 \times 10^{15} \text{ s}^{-1}} \times \frac{1 \times 10^9 \text{ nm}}{1 \text{ m}} = 102.5 \text{ nm}; \text{ ultraviolet.}$$

Yes, in the transition from a low level to a high one, energy is absorbed.

12. Refer to Section 6.2.

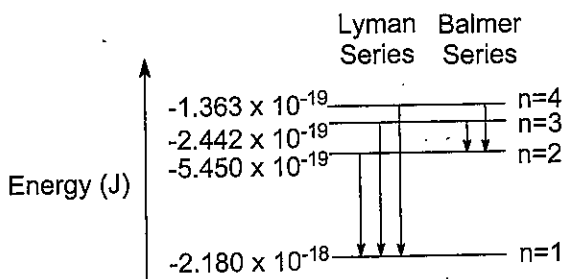
$$E_n = \frac{-R_H}{n^2}$$

$$E_1 = \frac{-R_H}{n^2} = \frac{2.180 \times 10^{-18} \text{ J}}{1^2} = -2.180 \times 10^{-18} \text{ J}$$

$$E_2 = \frac{-R_H}{n^2} = \frac{2.180 \times 10^{-18} \text{ J}}{2^2} = -5.450 \times 10^{-19} \text{ J}$$

$$E_3 = \frac{-R_H}{n^2} = \frac{2.180 \times 10^{-18} \text{ J}}{3^2} = -2.422 \times 10^{-19} \text{ J}$$

$$E_4 = \frac{-R_H}{n^2} = \frac{2.180 \times 10^{-18} \text{ J}}{4^2} = -1.363 \times 10^{-19} \text{ J}$$



Obviously, this graph assumes that transitions are from E_{hi} to E_{lo} . If one assumes the transitions are in the opposite direction, the only change would be in the directions of the arrows.

14. Refer to Section 6.2, Example 6.3, and Figure 6.2.

- (a) Use the Rydberg equation to calculate the frequency. Then calculate the wavelength to determine the region of the spectrum associated with that frequency.

For the transition $n = 5 \rightarrow n = 4$

$$\nu = \frac{R_H}{h} \left[\frac{1}{(n_{lo})^2} - \frac{1}{(n_{hi})^2} \right]$$

$$\nu = \frac{2.180 \times 10^{-18} \text{ J}}{6.626 \times 10^{-34} \text{ J} \cdot \text{s}} \left[\frac{1}{(4)^2} - \frac{1}{(5)^2} \right]$$

$$\nu = 7.403 \times 10^{13} / \text{s}$$

$$\lambda = \frac{c}{\nu} = \frac{2.998 \times 10^8 \text{ m/s}}{7.403 \times 10^{13} / \text{s}} \times \frac{10^9 \text{ nm}}{1 \text{ m}} = 4050 \text{ nm}$$

- (b) This wavelength of light occurs in the **infrared** region of the spectrum.

16. Refer to Section 6.2.

The longest wavelength corresponds to the smallest energy difference

$$\nu = \frac{R_H}{h} \left[\frac{1}{(n_{lo})^2} - \frac{1}{(n_{hi})^2} \right] = \frac{2.180 \times 10^{-18} \text{ J}}{6.626 \times 10^{-34} \text{ J} \cdot \text{s}} \left[\frac{1}{3^2} - \frac{1}{4^2} \right] = 1.599 \times 10^{14} \text{ s}^{-1}$$

$$\lambda = \frac{c}{\nu} = \frac{2.998 \times 10^8 \text{ m/s}}{1.599 \times 10^{14} / \text{s}} \times \frac{10^9 \text{ nm}}{1 \text{ m}} = 1875 \text{ nm}$$

18. Refer to Section 6.3.

Values of m_ℓ vary from $-\ell$ to $+\ell$ for any given ℓ value.

(a) p-sublevel: $\ell = 1$

$$m_\ell = -1, 0, +1.$$

(b) f-sublevel: $\ell = 3$

$$m_\ell = -3, -2, -1, 0, +1, +2, +3.$$

(c) For the $n=3$ shell, there are 3 values for $\ell = 0, 1, 2$.

$$\ell = 0: \text{ s-sublevel: } m_\ell = 0.$$

$$\ell = 1: \text{ p-sublevel: } m_\ell = -1, 0, +1.$$

$$\ell = 2: \text{ d-sublevel: } m_\ell = -2, -1, 0, +1, +2.$$

20. Refer to Section 6.3 and Figure 6.8.

Look at Figure 6.8 to determine the order of filling. The orbital that fills last is higher in energy.

(a) 3s

(b) 4d

(c) 4f is higher for atoms with $Z < 58$ and the 6s orbital is higher for atoms with $Z > 57$.

(d) 2s

22. Refer to Section 6.3.

The type of orbital is determined by the ℓ value.

(a) $\ell = 1$, therefore: p-orbital.

(b) $\ell = 0$, therefore: s-orbital.

(c) $\ell = 2$, therefore: d-orbital

24. Refer to Section 6.3.

The number of orbitals is equal to $2\ell + 1$. If there is more than one sublevel, add up the orbitals for each sublevel.

- (a) $n = 3$ has 3 sublevels, $\ell = 0$, $\ell = 1$ and $\ell = 2$
 $\ell = 0$: 1 orbital
 $\ell = 1$: 3 orbitals
 $\ell = 2$: 5 orbitals, thus the $n = 3$ shell has **9 orbitals**.
- (b) 4p: $\ell = 1$: **3 orbitals**
- (c) f: $\ell = 3$: **7 orbitals**
- (d) d: $\ell = 2$: **5 orbitals**

26. Refer to Section 6.3, Example 6.5, and Table 6.3.

- (a) These quantum numbers indicate a 1s orbital. The maximum number of electrons in a 1s orbital can be 2 electrons.
- (b) These quantum numbers indicate one of the 5f orbitals. The maximum number of electrons that can be in a single 5f orbital is 2 electrons.
- (c) These quantum numbers indicate a 3d orbital. Since there are 5 possible 3d orbitals and each can hold a maximum of 2 electrons, the total number of electrons that can reside in orbitals with these quantum numbers is 10 electrons.

28. Refer to Section 6.3.

Refer to the rules for quantum numbers. If a rule is violated, the set cannot occur.

(a) $n = 1 \Rightarrow \ell = 0$

$\ell = 0 \Rightarrow m_\ell = 0$

$m_\ell = 0 \Rightarrow m_s = -\frac{1}{2}, +\frac{1}{2}$

None of the rules are violated, so this set **can occur**.

(b) $n = 1 \Rightarrow \ell = 0, 1$

$\ell = 2$ This is not possible given the allowed values, this set **cannot occur**.

(c) $n = 3 \Rightarrow \ell = 0, 1, 2$

$\ell = 2 \Rightarrow m_\ell = -2, -1, 0, +1, +2$

$m_\ell = -2 \Rightarrow m_s = -\frac{1}{2}, +\frac{1}{2}$

None of the rules are violated, so this set **can occur**.

(d) $n = 2 \Rightarrow \ell = 0, 1$

$\ell = 1 \Rightarrow m_\ell = -1, 0, +1$

$m_\ell = -2$ This is not possible given the allowed values, this set **cannot occur**.

(e) $n = 4 \Rightarrow \ell = 0, 1, 2, 3$

$\ell = 0 \Rightarrow m_\ell = 0$

$m_\ell = 2$ This is not possible given the allowed values, this set **cannot occur**.

30. Refer to Section 6.5 and Examples 6.6 and 6.7.

Determine the number of electrons. Then write down the filling order and start filling the orbitals until you run out of electrons.

(a) B ($5e^-$): $1s^2 2s^2 2p^1$

(b) Ba ($56e^-$): $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2$

(c) Be ($4e^-$): $1s^2 2s^2$

(d) Bi ($83e^-$): $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 6p^3$

(e) Br ($35e^-$): $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^5$

32. Refer to Section 6.4, Examples 6.6 and 6.7, Table 6.4, and Figure 6.9.

Locate the element of interest on the periodic table. Move up one period and then move to the extreme right to find the appropriate noble gas. Which period is the element of interest in? That is the s-orbital from which you are to start filling. Then fill until you run out of electrons.

(a) Os, $[\text{Xe}] 6s^2 4f^{14} 5d^6$

(b) Mg, $[\text{Ne}] 3s^2$

(c) Ge, $[\text{Ar}] 4s^2 3d^{10} 4p^2$

(d) V, $[\text{Ar}] 4s^2 3d^3$

(e) At, $[\text{Xe}] 6s^2 4f^{14} 5d^{10} 6p^5$

34. Refer to Section 6.5.

Start writing an electron configuration, filling up the orbitals until you meet the criterion given. Then determine the element that configuration corresponds to.

(a) $[\text{Xe}] 6s^2 4f^{14}$

This corresponds to **Yb**.

(b) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^2$
This corresponds to **Sn**.

(c) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^2$
This corresponds to **Zr**.

(d) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^5$
This corresponds to **I**.

36. Refer to Section 6.5.

Write out the electron configuration, count up the number of electrons in p subshells and divide that number by the total number of electrons.

(a) Mg: $1s^2 2s^2 2p^6 3s^2$
(6 p electrons) / (12 total electrons) = **0.5** (or one half)

(b) Mn: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^5$
(12 p electrons) / (25 total electrons) = **0.48** (or almost one half)

(c) Mo: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^4$
(18 p electrons) / (42 total electrons) = **0.43** (little more than two fifths)

38. Refer to Section 6.5, Figure 6.8, and Example 6.6.

First determine if the given configuration violates one of the rules for quantum numbers. If it does not, determine the ground state configuration for an atom with the given number of electrons and compare that with the configuration given.

(a) 3 electrons, ground state: $1s^2 2s^1$
Thus the given configuration is an **excited state**.

(b) 8 electrons, ground state: $1s^2 2s^2 2p^4$
Thus the given configuration is the **ground state**.

(c) 10 electrons, ground state: $1s^2 2s^2 2p^6$
Thus the given configuration is the **excited state**.

(d) The p orbital can have a maximum of 6 electrons.
Impossible configuration.

(e) The d orbital can have a maximum of 10 electrons.
Impossible configuration.

40. Refer to Section 6.6 and Example 6.8.

- (a) N
a: 1s 2s 2p 3s
 (↑↓) (↑↓) (↑↓)(↑↓)(↑↓) (↑)
- (b) O: 1s 2s 2p
 (↑↓) (↑↓) (↑↓)(↑)(↑)
- (c) C 1s 2s 2p 3s 3p 4s 3d
o: (↑↓) (↑↓) (↑↓)(↑↓)(↑↓) (↑↓) (↑↓)(↑↓)(↑↓) (↑↓) (↑↓)(↑↓)(↑)(↑)(↑)
- (d) Cl 1s 2s 2p 3s 3p
 (↑↓) (↑↓) (↑↓)(↑↓)(↑↓) (↑↓) (↑↓)(↑↓)(↑)

42. Refer to Section 6.6.

Count the number of electrons. Since the number of electrons must equal the number of protons (atomic number) in a neutral atom, you can then identify the element from the periodic table.

- (a) 12 electrons, $Z = 12$, therefore: **Mg**.
 (b) 15 electrons, $Z = 15$, therefore: **P**.
 (c) 8 electrons, $Z = 8$, therefore: **O**.

44. Refer to Section 6.6 and Figure 6.9.

- (a) Sn has 2 half-filled 5p orbitals, Sb has 3 half-filled 5p orbitals, and Te has 2 half-filled 5p orbitals.
 (b) K, Rb, Cs, Fr; every Group 1 element past Ar which has filled 3p orbitals.
 (c) Ge, As, Sb, Te; every metalloid past S that has paired 3p electrons.
 (d) None, the only element that has filled 3d orbitals and 3 half-filled 3p orbitals is As, but As is a metalloid, not a nonmetal.

46. Refer to Section 6.6.

All the inner sublevels will be filled, so only the outermost sublevel will contain unpaired electrons.

- (a) Hg: $[\text{Xe}]6s^2 4f^{14} 5d^{10}$ $5d^{10}$: $(\uparrow\downarrow)(\uparrow\downarrow)(\uparrow\downarrow)(\uparrow\downarrow)(\uparrow\downarrow)$ **0 unpaired electrons.**
- (b) Mn: $[\text{Ar}]4s^2 3d^5$ $3d^5$: $(\uparrow)(\uparrow)(\uparrow)(\uparrow)(\uparrow)$ **5 unpaired electrons.**
- (c) Mg: $[\text{Ne}]3s^2$ $3s^2$: $(\uparrow\downarrow)$ **0 unpaired electrons.**

48. Refer to Section 6.6.

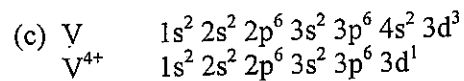
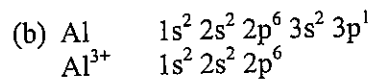
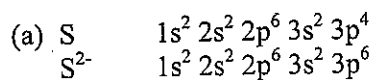
The only main group metals in the 4th period are K, Ca and Ga, the remainder of the elements are transition metals, nonmetals or metalloids.

	4s	4p		
K	(\uparrow)	$()$	$()$	$()$
Ca	$(\uparrow\downarrow)$	$()$	$()$	$()$
Ga	$(\uparrow\downarrow)$	(\uparrow)	$()$	$()$

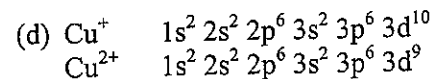
- (a) Ca
 (b) K and Ga
 (c) none
 (d) none

50. Refer to Section 6.7 and Example 6.9.

Write the ground state electron configuration for the atom, then remove (for cations) or add (for anions) the number of electrons indicated by the charge.



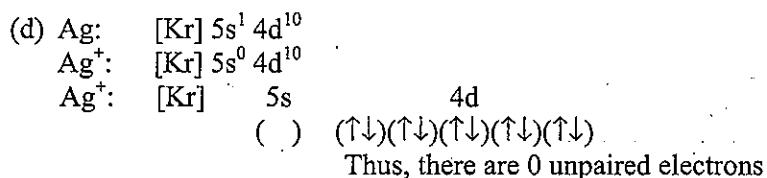
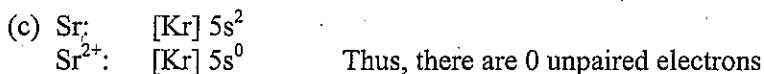
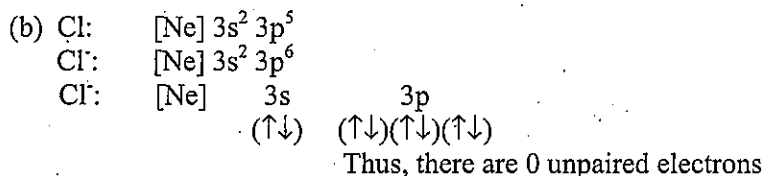
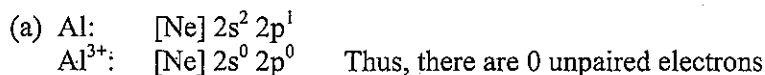
For transition metals, electrons are lost first from the outermost s sublevel.



Remember that the electron configuration of Cu is $[\text{Ar}] 4s^1 3d^{10}$ and that for transition metals, electrons are lost first from the outermost s sublevel.

52. Refer to Section 6.7.

Write the ground state electron configuration for the atom, then remove (for cations) or add (for anions) the number of electrons indicated by the charge. Then write the orbital diagram for the ion.



54. Refer to Section 6.8, Figures 6.13 and 6.15, and Example 6.11.

- (a) Atomic radius increases from right to left across a period, therefore:
 $S < Si < Na$
- (b) Ionization energy increases from left to right across a period, therefore:
 $Na < Si < S$
- (c) Electronegativity decreases from right to left across a period, therefore:
 $S > Si > Na$

56. Refer to Section 6.8 and Example 6.10.

- (a) Atomic radius decreases from left to right across a period and increases down a group, thus the largest atom would be the lower leftmost: **K**.
- (b) Ionization energy increases from left to right across a period and decreases down a group, thus the atom with the highest ionization energy would be the upper rightmost: **Cl**.
- (c) Electronegativity increases from left to right across a period and decreases down a group, thus the most electronegative atom would be the upper rightmost: **Cl**.

58. Refer to Section 6.8 and Example 6.10.

Cations are smaller than the corresponding atoms, anions are larger.

- (a) N
- (b) Ba⁺²
- (c) Se
- (d) Co⁺³

60. Refer to Section 6.8, Figure 6.13, and Example 6.10.

Atomic radii decrease from left to right across a period and increases down a group. Thus, the ordering of radii from smallest to largest is:

- (a) Kr < K < Rb < Cs
- (b) Ar < Si < Al < Cs

62. Refer to Section 6.1 and Figure 6.3.

(a) Both wavelengths fall into the green region of visible light.

$$(b) 486 \text{ nm} \times \frac{1 \text{ m}}{1 \times 10^9 \text{ nm}} = 4.85 \times 10^{-7} \text{ m}$$

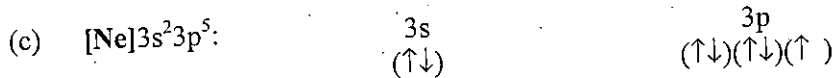
$$E = \frac{hc}{\lambda} = \frac{(6.626 \times 10^{-34} \text{ J}\cdot\text{s})(2.998 \times 10^8 \text{ m/s})}{4.85 \times 10^{-7} \text{ m}} = 4.10 \times 10^{-19} \text{ J}$$

$$\frac{4.10 \times 10^{-19} \text{ J}}{1 \text{ photon}} \times \frac{6.02 \times 10^{23} \text{ photons}}{1 \text{ mol.}} = 2.47 \times 10^5 \text{ J/mol.} = 2.47 \times 10^2 \text{ kJ/mol.}$$

$$512 \text{ nm} \times \frac{1 \text{ m}}{1 \times 10^9 \text{ nm}} = 5.12 \times 10^{-7} \text{ m}$$

$$E = \frac{hc}{\lambda} = \frac{(6.626 \times 10^{-34} \text{ J}\cdot\text{s})(2.998 \times 10^8 \text{ m/s})}{5.12 \times 10^{-7} \text{ m}} = 3.88 \times 10^{-19} \text{ J}$$

$$\frac{3.88 \times 10^{-19} \text{ J}}{1 \text{ photon}} \times \frac{6.02 \times 10^{23} \text{ photons}}{1 \text{ mol.}} = 2.34 \times 10^5 \text{ J/mol.} = 2.34 \times 10^2 \text{ kJ/mol.}$$



64. Refer to Sections 6.5, 6.7, and 6.8.

- (a) 17 electrons, therefore 17 protons, thus the element is **Cl**.
- (b) Since ionization energy decreases going down a group, the element is **Pb**.
- (c) The +2 ion has 23 electrons, therefore the parent atom has 25 and $Z = 25$, which is **Mn**.
- (d) Since atomic radii increase going down a group, the smallest would be **Li**.
- (e) Ionization energy increases from left to right across a period, so the element with the greatest ionization energy is **Kr**.

66. Refer to Section 6.2.

- (a) Energy is absorbed when the electrons are excited from lower levels to a higher ones, thus energy is absorbed for **transitions 2 and 4**.
- (b) Energy is emitted when the electron relaxes from higher level to a lower one, thus energy is emitted for **transitions 1 and 3**.
- (c) We cannot answer this question without knowing what the element is since one must know the number of electrons to determine the ground state configuration. If we assume the atom to be hydrogen, then **transition 1** involves the ground state.
- (d) The transition with largest energy difference (see part (a) and problem 12 above) will absorb the most energy, thus **transition 2**.
- (e) The transition with largest energy difference (see part (b) and problem 12 above) will emit the most energy, thus **transition 1**.

68. Refer to Section 6.5.

Locate the element of interest on the periodic table. Move up one period and then move to the extreme right to find the appropriate noble gas. Which period is the element of interest in? That is the s-orbital from which you are to start filling. Then fill until you run out of electrons.

- (a) Li $[\text{He}] 2s^1$
- (b) Ra $[\text{Rn}] 7s^2$
- (c) Sc $[\text{Ar}] 4s^2 3d^1$
- (d) Sb $[\text{Kr}] 5s^2 4d^{10} 5p^3$
- (e) Ca $[\text{Ar}] 4s^2$

70. Refer to Section 6.2 and Figures 6.1 and 6.7.

- (a) There are two main differences between the Bohr model and the quantum mechanical model.
1. The kinetic energy of an electron is inversely related to the space it occupies.
 2. You cannot know the position of an electron at any given instant.
- (b) Frequency and wavelength are inversely related. If frequency is short, wavelength is long and vice versa.
- (c) The three p orbitals lie along a different axis, x, y, or z. They are all shaped the same, like two small balloons with the knot ends joined. This shape is often referred to as dumbbell shaped.

72. Refer to Sections 6.1 – 6.5 and Example 6.7.

- (a) This statement is **true**. Photons with very short wavelengths are high energy.
- (b) This statement is **false**. The energy of an electron is inversely proportional to n^2 , not l .

$$E_n = -\frac{R_H}{n^2}$$

- (c) This statement is **false**. Electrons start entering the 5th principal level **before** the fourth is filled. The order of filling is [Kr] 5s 4d 5p ...

74. Refer to Sections 6.6 and 6.7.

- (a) True
- (b) True
- (c) False Energy is absorbed when an electron is removed from an atom.

76. Refer to Section 6.2.

For a one electron species, the ground state is $n = 1$, so the first excited state is $n = 2$.

$$E = \frac{-BZ^2}{n^2} = \frac{-(2.180 \times 10^{-18} \text{ J})(3)^2}{2^2} = -4.905 \times 10^{-18} \text{ J}$$

$$E = -4.905 \times 10^{-18} \text{ J} \times \frac{1 \text{ kJ}}{1000 \text{ J}} \times \frac{6.022 \times 10^{23}}{1 \text{ mol}} = -2.954 \times 10^3 \text{ kJ/mol.}$$

E is the energy of the electron. The energy needed to ionize (or remove) an electron is $+2.954 \times 10^3 \text{ kJ/mol}$.

77. Refer to Sections 6.1 and 6.2.

$$\lambda = \frac{hc}{\Delta E} \quad \text{and} \quad \Delta E = -R_H \left[\frac{1}{n_{hi}^2} - \frac{1}{n_{lo}^2} \right]$$

Substituting for ΔE and $n_{lo} = 2$, we get: $\lambda = \frac{hc}{-R_H \left[\frac{1}{n_{hi}^2} - \frac{1}{2^2} \right]} = -\frac{hc}{R_H} \frac{1}{\left[\frac{1}{n_{hi}^2} - \frac{1}{4} \right]}$

$$\lambda = \frac{(6.626 \times 10^{-34} \text{ J} \cdot \text{s})(2.998 \times 10^8 \text{ m/s})}{2.180 \times 10^{-18} \text{ J}} \frac{1}{\left[\frac{1}{n_{hi}^2} - \frac{1}{4} \right]} = -9.112 \times 10^{-8} \text{ m} \frac{1}{\left[\frac{1}{n_{hi}^2} - \frac{1}{4} \right]}$$

$$\lambda = 9.112 \times 10^{-8} \text{ m} \frac{1}{\left[\frac{1}{4} - \frac{1}{n_{hi}^2} \right]} = 9.112 \times 10^{-8} \text{ m} \frac{4n_{hi}^2}{\frac{4n_{hi}^2}{4} - \frac{4n_{hi}^2}{n_{hi}^2}}$$

$$\lambda = 9.112 \times 10^{-8} \text{ m} \frac{4n_{hi}^2}{n_{hi}^2 - 4} = \frac{(3.645 \times 10^{-7} \text{ m})n_{hi}^2}{n_{hi}^2 - 4}$$

Converting to nanometers: $\lambda = \frac{(3.645 \times 10^2 \text{ nm})n_{hi}^2}{n_{hi}^2 - 4}$

78. Refer to Section 6.3 and Table 6.3.

n	1			2				
l	0	1		0	1	2		
sublevel	1s	1p		2s	2p		2d	
m _l	0 1	0 1 2		0 1	0 1 2		0 1 2 3	

electron configuration with eight electrons: $1s^4 1p^4$

79. Refer to Section 6.3 and Table 6.3.

- (a) If there were 3 values for m_s , then each orbital could hold 3 electrons.
 s-sublevel: one orbital ($m_l = 0$), 3 electrons.
 p-sublevel: 3 orbitals ($m_l = -1, 0, 1$), 9 electrons.
 d-sublevel: 5 orbitals ($m_l = -2, -1, 0, 1, 2$), 15 electrons.
- (b) The $n = 3$ level could hold 27 electrons; 3 in the s-sublevel, 9 in the p-sublevel and 15 in the d-sublevel.
- (c) AN = 8: $1s^3 2s^3 2p^2$
 AN = 17: $1s^3 2s^3 2p^9 3s^2$

80. Refer to Section 6.1.

Calculate the energy of the light. The difference between that energy and the kinetic energy is the energy needed to eject the electron (E_{\min}).

$$(a) E = \frac{hc}{\lambda} = \frac{(6.626 \times 10^{-34} \text{ J} \cdot \text{s})(2.998 \times 10^8 \text{ m/s})}{5.40 \times 10^{-7} \text{ m}} = 3.68 \times 10^{-19} \text{ J}$$

$$E_{\min} = 3.68 \times 10^{-19} \text{ J} - 2.60 \times 10^{-20} \text{ J} = 3.42 \times 10^{-19} \text{ J}$$

While the calculations above used the longer wavelength, we also could have used the shorter wavelength and gotten the same answer (within significance).

$$E = \frac{hc}{\lambda} = \frac{(6.626 \times 10^{-34} \text{ J} \cdot \text{s})(2.998 \times 10^8 \text{ m/s})}{4.00 \times 10^{-7} \text{ m}} = 4.97 \times 10^{-19} \text{ J}$$

$$E_{\min} = 4.97 \times 10^{-19} \text{ J} - 1.54 \times 10^{-19} \text{ J} = 3.43 \times 10^{-19} \text{ J}$$

- (b) Since longer wavelengths correspond to lower energies, E_{\min} would correspond to the longest wavelength.

$$\lambda = \frac{hc}{E_{\min}} = \frac{(6.626 \times 10^{-34} \text{ J} \cdot \text{s})(2.998 \times 10^8 \text{ m/s})}{3.42 \times 10^{-19} \text{ J}} = 5.81 \times 10^{-7} \text{ m} = 581 \text{ nm}$$