Text Problems: 6, 8, 10, 18, 20, 24, 26, 28, 30, 34, 38, 42, 44, 50, 54, 56, 58

6. Refer to Example 6.1.

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

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

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.

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

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

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

18. Refer to Section 6.3.

Values of $\mathbf{m}_\ell$ vary from $-\ell$ to $+\ell$ for any given ℓ value.

a. p-sublevel: $\ell = 1$
$\mathbf{m}_\ell = -1, 0, +1.$

b. f-sublevel: $\ell = 3$
$\mathbf{m}_\ell = -3, -2, -1, 0, +1, +2, +3.$

c. For the $\mathbf{n}$=3 shell, there are 3 values for $\ell = 0, 1, 2.$
$\ell = 0$: s-sublevel: $\mathbf{m}_\ell = 0.$
$\ell = 1$: p-sublevel: $\mathbf{m}_\ell = -1, 0, +1.$
$\ell = 2$: d-sublevel: $\mathbf{m}_\ell = -2, -1, 0, +1, +2.$

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

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. $\mathbf{n} = 3$ has 3 sublevels, $\ell = 0$, $\ell = 1$ and $\ell = 2$
$\ell = 0$: 1 orbital
$\ell = 1$: 3 orbitals
$\ell = 2$: 5 orbitals, thus the $\mathbf{n} = 3$ 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. $\mathbf{n} = 1 \Rightarrow \ell = 0$
$\ell = 0 \Rightarrow \mathbf{m}_\ell = 0$
$\mathbf{m}_\ell = 0 \Rightarrow \mathbf{m}_s = -\frac{1}{2}, +\frac{1}{2}$
None of the rules are violated, so this set **can occur.**

b. $\mathbf{n} = 1 \Rightarrow \ell = 0, 1$
$\ell = 2$ This is not possible given the allowed values, this set **cannot occur.**

c. $\mathbf{n} = 3 \Rightarrow \ell = 0, 1, 2$
$\ell = 2 \Rightarrow \mathbf{m}_\ell = -2, -1, 0, +1, +2$
$\mathbf{m}_\ell = -2 \Rightarrow \mathbf{m}_s = -\frac{1}{2}, +\frac{1}{2}$
None of the rules are violated, so this set **can occur.**

d. $\mathbf{n} = 2 \Rightarrow \ell = 0, 1$
$\ell = 1 \Rightarrow \mathbf{m}_\ell = -1, 0, +1$
$\mathbf{m}_\ell = -2$ This is not possible given the allowed values, this set **cannot occur.**

e. $\mathbf{n} = 4 \Rightarrow \ell = 0, 1, 2, 3$
$\ell = 0 \Rightarrow \mathbf{m}_\ell = 0$
$\mathbf{m}_\ell = 2$ 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^-): $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^5$

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. [Xe]$6s^2\ 4f^{14}$
This corresponds to **Yb.**

This is an unusual circumstance and you would not be expected to memorize it.

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.**

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.

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

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.

a. S $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$
   S^{2-} $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

b. Al $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$
   Al^{3+} $1s^2\ 2s^2\ 2p^6$

c. V $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^3$
   V^{4+} $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^1$
   For transition metals, electrons are lost first from the outermost s sublevel.

d. Cu^+ $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}$
   Cu^{2+} $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^9$
   Remember that the electron configuration of Cu is [Ar] $4s^1\ 3d^{10}$ *and that for transition metals, electrons are lost first from the outermost s sublevel.*

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}