Problems of the Day up to 11/5/09

The following answers are cut and pasted from the publisher's answer keys. Any additional commentary by me is in {} brackets.

4.70 First calculate the final volume of the dilute solution. Then, you can subtract 505 mL from this volume to calculate the amount of water that should be added.

$$V_d = \frac{M_d \times V_d}{M_c} = \frac{0.125\,M \times 505\text{ mL}}{0.100\,M} = 631\text{ mL}$$

(631 − 505) mL = **126 mL of water**

6.58 The allowed values of l are 0, 1, 2, 3, and 4. These correspond to the 5*s*, 5*p*, 5*d*, 5*f*, and 5*g* subshells. These subshells each have one, three, five, seven, and nine orbitals, respectively.

{**6.59** *s*: spherical; *p*: "dumbbell"; *d*: four "cloverleaf", one "dumbbell with hula hoop". l determines whether they are *s*, *p*, *d*, etc.; n limits the values of l that are allowed and describes their relative energy (e.g. *4d* are higher in energy than *3d*); the number of values of m_l allowed translates to the number of degenerate orbitals there are of that type (one *s*, three *p*, 5 *d*, etc. per n level).

6.64 **(a)** The number given in the designation of the subshell is the principal quantum number, so in this case $\boldsymbol{n = 3}$. For *s* orbitals, $\boldsymbol{l = 0}$. m_l can have integer values from $-l$ to $+l$, therefore, $\boldsymbol{m_l = 0}$. The electron spin quantum number, $\boldsymbol{m_s}$, can be either **+1/2** or **−1/2**.

Following the same reasoning as part **(a)**

(b) 4*p*: $n = 4$; $l = 1$; $m_l = -1, 0, 1$; $m_s = +1/2, -1/2$

(c) 3*d*: $n = 3$; $l = 2$; $m_l = -2, -1, 0, 1, 2$; $m_s = +1/2, -1/2$

6.76 3*s*: two 3*d*: ten 4*p*: six 4*f*: fourteen 5*f*: fourteen

6.77 The electron configurations for the elements are

(a) N: $1s^22s^22p^3$ There are three *p*-type electrons.

(b) Si: $1s^22s^22p^63s^23p^2$ There are six *s*-type electrons.

(c) S: $1s^22s^22p^63s^23p^4$ There are no *d*-type electrons.

6.78 **(a)** is wrong because the magnetic quantum number m_l can have only whole number values.

(b) is wrong because, when the angular momentum quantum number l is 0, the magnetic quantum number m_l can only have the value 0.

(c) same as (b)

(d) **OK**

(e) is wrong because the electron spin quantum number m_s can have only half-integral values.

6.79 For aluminum, there are two few 2*p* electrons. (The 2*p* subshell holds six electrons.) The number of electrons (13) is correct. The electron configuration should be $1s^22s^22p^63s^23p^1$. The configuration shown might be an excited state of an aluminum atom.

For boron, there are too many $2p$ electrons. (Boron only has five electrons.) The electron configuration should be $1s^22s^22p^1$. What would be the electric charge of a boron ion with the electron arrangement given in the problem?

For fluorine, there are also too many $2p$ electrons. (Fluorine only has nine electrons.) The configuration shown is that of the F^- ion. The correct electron configuration is $1s^22s^22p^5$.

6.93 **Strategy:** How many electrons are in the Ge atom (Z = 32)? We start with $n = 1$ and proceed to fill orbitals in the order shown in Figure 6.23 of the text. Remember that any given orbital can hold at most 2 electrons. However, don't forget about degenerate orbitals. Starting with $n = 2$, there are three p orbitals of equal energy, corresponding to $m_l = -1, 0, 1$. Starting with $n = 3$, there are five d orbitals of equal energy, corresponding to $m_l = -2, -1, 0, 1, 2$. We can place electrons in the orbitals according to the Pauli exclusion principle and Hund's rule. The task is simplified if we use the noble gas core preceding Ge for the inner electrons.

Solution: Germanium has 32 electrons. The noble gas core in this case is [Ar]. (Ar is the noble gas in the period preceding germanium.) [Ar] represents $1s^22s^22p^63s^23p^6$. This core accounts for 18 electrons, which leaves 14 electrons to place.

See Figure 6.23 of your text to check the order of filling subshells past the Ar noble gas core. You should find that the order of filling is $4s$, $3d$, $4p$. There are 14 remaining electrons to distribute among these orbitals. The $4s$ orbital can hold two electrons. Each of the five $3d$ orbitals can hold two electrons for a total of *10* electrons. This leaves two electrons to place in the $4p$ orbitals.

The electrons configuration for Ge is:

$$[Ar]4s^23d^{10}4p^2$$

You should follow the same reasoning for the remaining atoms.

Fe: $[Ar]4s^23d^6$	Zn: $[Ar]4s^23d^{10}$	Ni: $[Ar]4s^23d^8$
W: $[Xe]6s^24f^{14}5d^4$	Tl: $[Xe]6s^24f^{14}5d^{10}6p^1$	

{7.27 p.246
7.28 p.247-249
7.30 p.250-251. }

7.68 The ions listed are all isoelectronic. The ion with the fewest protons will have the largest ionic radius, and the ion with the most protons will have the smallest ionic radius. The order of increasing atomic radius is:

$$\mathbf{Mg^{2+} < Na^+ < F^- < O^{2-} < N^{3-}}$$

{7.72 They have the same valence electronic configuration, i.e. their valence electrons are the same in number and in the orbitals they occupy. It’s the outermost valence electrons that are involved in chemical reactions.}