Please show your work for all calculations, and report answers to the proper number of significant digits to receive full credit. For calculations, circle your final answer.

1. A certain type of electromagnetic radiation has a frequency of $3.00 \times 10^{14}$ Hz.
   a. Determine the wavelength ($\lambda$) of the radiation.

   $\lambda = \frac{c}{\nu} = \frac{2.998 \times 10^{8} \text{ m/s}}{3.00 \times 10^{14} \text{ s}^{-1}} = 9.99 \times 10^{-7} \text{ m} = 999 \text{ nm}$

   b. Determine the energy of a photon of this radiation.

   $E = h\nu = (6.626 \times 10^{-34} \text{ Js})(3.00 \times 10^{14} \text{ s}^{-1}) = 1.99 \times 10^{-19} \text{ J}$

   c. Determine the mass of a photon of this radiation.

   $m = \frac{h}{\lambda c} = \frac{6.626 \times 10^{-34} \text{ Js}}{(9.99 \times 10^{-7} \text{ m})(2.998 \times 10^{8} \text{ m/s})} = 2.21 \times 10^{-36} \text{ kg}$

   d. In what region of the electromagnetic spectrum does this radiation fall? (Refer to Fig. 7.4.)

   infrared

2. In your own words, explain why hydrogen atoms produce line emission spectra instead of continuous emission spectra.

   Atomic emission spectra are produced by the photons released when electrons fall from higher to lower energy. Since the energy of the electron is quantized, only certain energy changes are possible, and these are the energies that show up in the line spectra.

   Continued on reverse.
3. Determine whether each of the following sets of quantum numbers is or is not allowed and explain why each not allowed combination is not possible.
   a. $n = 4; l = 2; m_l = -1; m_s = \frac{1}{2}$  \textit{Allowed}
   
   b. $n = 4; l = -2; m_l = -2; m_s = -\frac{1}{2}$ Not allowed. $l$ must be between 0 and $n-1$. It cannot be negative.
   
   c. $n = 2; l = 2; m_l = -1; m_s = \frac{1}{2}$ Not allowed. $l$ must be less than $n$.
   
   d. $n = 2; l = 0; m_l = 0; m_s = -\frac{1}{2}$ \textit{Allowed}.

4. Write the electron configuration (using the noble gas notation) for each of the following elements or ions.
   a. sulfur $16e^- \quad S: [Ne]3s^23p^4$
   
   b. Cl $18e^- \quad Cl^- : [Ar]$ 
   
   c. Mn $25e^- \quad Mn : [Ar]4s^23d^5$
   
   d. Mn$^{2+}$ $23e^- \quad Mn^{2+} : [Ar]3d^5$

5. By using only their locations in the periodic table (no numbers), put the following elements in order from smallest to largest ionization energy: I, Rb, O, C

   \text{Lowest} \quad \text{Rb} \quad \text{I} \quad \text{C} \quad \text{O} \quad \text{Highest}

   I$_i$ is highest at the upper right and lowest at the lower left of the periodic table.