

1. Show the electron configurations of all the electrons (then show the electron configurations of the highest-energy electron) using the quantum number technique for each of the following atoms.



Quantum #	$\frac{m_l}{l}$	$\frac{m_s}{s}$
	2	2



Quantum #	$\frac{m_l}{l}$	$\frac{m_s}{s}$
	3	2



Quantum #	$\frac{m_l}{l}$	$\frac{m_s}{s}$
	3	2

2. Calculate the energy needed to cause an electron to jump from the 3rd shell to the 2nd shell. Energy =

$$E = 2.18 \times 10^{-18} \text{ J} \left(\frac{1}{n_1^2} - \frac{1}{n_2^2} \right) = 2.18 \times 10^{-18} \text{ J} \left(\frac{1}{3^2} - \frac{1}{2^2} \right) = 2.18 \times 10^{-18} \text{ J} \left(\frac{1}{9} - \frac{1}{4} \right) = 2.18 \times 10^{-18} \text{ J} \left(-\frac{5}{36} \right) = 2.97 \times 10^{-19} \text{ J}$$

3. Calculate the wavelength for the photon that will be emitted when an electron falls spontaneously to the 2nd shell of a hydrogen atom. Wavelength =

$$E = h\nu = \frac{hc}{\lambda} \quad \nu = \frac{c}{\lambda} \quad \text{from these we get}$$

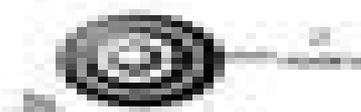
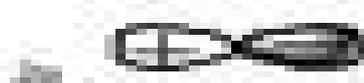
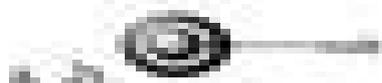
$$\lambda = \frac{hc}{E} = \frac{6.626 \times 10^{-34} \text{ J}\cdot\text{s} \times 3.00 \times 10^8 \text{ m/s}}{2.97 \times 10^{-19} \text{ J}} = 6.72 \times 10^{-7} \text{ m}$$

4. What does the solution to the Schrodinger equation for an electron represent?

The solution to the Schrodinger equation represents the energy of an electron in an atomic orbital described by the wavefunction ψ .

$$\hat{H}\psi = E\psi$$

5. Draw the shape of the following orbitals.



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