7.12 The energy of one photon of light is described by the equation, $E = h\nu = \frac{hc}{\lambda}$. To obtain energy per mol of photons, multiply by the Avogadro constant. Begin by converting the wavelength into meters and then determine the energy.

(a) $\lambda = 665.7 \text{ nm}\left(\frac{10^{-9}\text{ m}}{1\text{ nm}}\right) = 6.657 \times 10^{-7}\text{ m}$

$$E_{\text{photon}} = \frac{(6.626 \times 10^{-34}\text{ J s})(2.998 \times 10^{8}\text{ m/s})}{6.657 \times 10^{-7}\text{ m}} = 2.984 \times 10^{-19}\text{ J/photon;}$$

$$E_{\text{molar}} = 2.984 \times 10^{-19}\text{ J}\left(\frac{6.022 \times 10^{23}\text{ photons}}{1\text{ mol}}\right)\left(\frac{10^{-3}\text{ kJ}}{1\text{ J}}\right) = 1.797 \times 10^{2}\text{ kJ/mol;}$$

(b) $\lambda = 1255 \text{ nm}\left(\frac{10^{-9}\text{ m}}{1\text{ nm}}\right) = 1.255 \times 10^{-6}\text{ m}$

$$E_{\text{photon}} = \frac{(6.626 \times 10^{-34}\text{ J s})(2.998 \times 10^{8}\text{ m/s})}{1.255 \times 10^{-6}\text{ m}} = 1.583 \times 10^{-19}\text{ J/photon;}$$

$$E_{\text{molar}} = 1.589 \times 10^{-19}\text{ J}\left(\frac{6.022 \times 10^{23}\text{ photons}}{1\text{ mol}}\right)\left(\frac{10^{-3}\text{ kJ}}{1\text{ J}}\right) = 9.569 \times 10^{1}\text{ kJ/mol;}$$

(c) $E_{\text{photon}} = (6.626 \times 10^{-34}\text{ J s})(4.5528 \times 10^{15}\text{ s}^{-1}) = 3.0167 \times 10^{-18}\text{ J/photon;}$

$$E_{\text{molar}} = 3.0167 \times 10^{-18}\text{ J}\left(\frac{6.022 \times 10^{23}\text{ photons}}{1\text{ mol}}\right)\left(\frac{10^{-3}\text{ kJ}}{1\text{ J}}\right) = 1.8167 \times 10^{3}\text{ kJ/mol.}$$

7.14 First, calculate the energy of one photon; then use $E_{\text{total}} = n\,E_{\text{photon}}$ to determine E_{total}:

$$\lambda = 450 \text{ nm}\left(\frac{10^{-9}\text{ m}}{1\text{ nm}}\right) = 4.50 \times 10^{-7}\text{ m;}$$

$$E_{\text{photon}} = \frac{(6.626 \times 10^{-34}\text{ J s})(2.998 \times 10^{8}\text{ m/s})}{4.50 \times 10^{-7}\text{ m}} = 4.41 \times 10^{-19}\text{ J/photon;}$$

$$E_{\text{total}} = 2.75 \times 10^{15}\text{ photons}\left(\frac{4.41 \times 10^{-19}\text{ J}}{1\text{ photon}}\right)\left(\frac{10^{3}\text{ mJ}}{1\text{ J}}\right) = 1.21\text{ mJ}$$

7.16 (a) Convert from molar energy to energy per photon by dividing by the Avogadro constant, then use $E = \frac{hc}{\lambda}$ to find $\lambda = \frac{hc}{E}$.

$$E_{\text{photon}} = \left(\frac{355\text{ J}}{1\text{ mol}}\right)\left(\frac{1\text{ mol}}{6.022 \times 10^{23}\text{ photons}}\right) = 5.90 \times 10^{-22}\text{ J;}$$

$$\lambda = \frac{(6.626 \text{ x } 10^{-34} \text{ J s})(2.998 \text{ x } 10^{8} \text{ m / s})}{5.90 \text{ x } 10^{-22} \text{ J}} = 3.37 \text{ x } 10^{-4} \text{ m}$$

$$\nu = \frac{c}{\lambda} = \frac{2.998 \text{x} 10^{8} \text{ m / s}}{3.37 \text{ x } 10^{-4} \text{m}} = 8.90 \text{ x } 10^{11} \text{ Hz}$$

(b) Use $E = \frac{hc}{\lambda}$ to find $\lambda = \frac{hc}{E}$.

$E_{\text{photon}} = 2.50 \text{ x } 10^{-18}$ J;

$$\lambda = \frac{(6.626 \text{ x } 10^{-34} \text{ J s})(2.998 \text{ x } 10^{8} \text{ m / s})}{2.50 \text{ x } 10^{-18} \text{ J}} = 7.95 \text{ x } 10^{-8} \text{ m or } 79.5 \text{ nm}$$

$$\nu = \frac{c}{\lambda} = \frac{2.998 \text{ x } 10^{8} \text{ m/s}}{7.95 \text{ x } 10^{-8} \text{ m}} = 3.77 \text{ x } 10^{15} \text{ Hz}$$

7.18 (a) The longest wavelength that will eject electrons corresponds to a photon with energy equal to the binding energy: $\lambda = \frac{hc}{bE}$.

$$\lambda = \frac{(6.626 \text{ x } 10^{-34} \text{ J s})(2.998 \text{ x } 10^{8} \text{ m / s})}{7.21 \text{ x } 10^{-19} \text{ J}} = 2.76 \text{ x } 10^{-7} \text{ m or } 276 \text{ nm.}$$

(b) First determine the energy of the photon, using $bE = E_{\text{photon}} - kE_{\text{electron}}$; then calculate the frequency from $E_{\text{photon}} = h\nu$.

$E_{\text{photon}} = bE + kE_{\text{electron}} = (7.21 \text{ x } 10^{-19} \text{ J}) + (2.5 \text{ x } 10^{-19} \text{ J}) = 9.7 \text{ x } 10^{-19}$ J;

$$\nu = \frac{E_{photon}}{h} = \frac{9.7 \text{ x } 10^{-19} \text{ J}}{6.626 \text{ x } 10^{-34} \text{ J s}} = 1.5 \text{ x } 10^{15} \text{ s}^{-1};$$

(c) $\lambda = \frac{c}{\nu} = \frac{2.998 \text{ x } 10^{8} \text{ m / s}}{1.5 \text{ x } 10^{15} \text{ s}^{-1}} = 2.0 \text{ x } 10^{-7} \text{ m or } 200 \text{ nm.}$

7.72 The photoelectric effect is described by an energy balance equation:

$E_{\text{kinetic, electron}} = E_{\text{photon}} - bE$, and $E = h\nu = \frac{hc}{\lambda}$

(a) $bE = h\nu = (6.626 \text{ x } 10^{-34} \text{ J s})(7.5 \text{ x } 10^{14} \text{ s}^{-1}) = 5.0 \text{ x } 10^{-19}$ J;

(b) $E_{\text{photon}} = \frac{(6.626 \text{x} 10^{-34} \text{ J s})(2.998 \text{ x } 10^{8} \text{ m / s})}{3.66 \text{ x } 10^{-7} \text{ nm}} = 5.43 \text{ x } 10^{-19}$ J;

$E_{\text{kinetic, electron}} = E_{\text{photon}} - bE = (5.43 \text{ x } 10^{-19} \text{ J}) - (5.0 \text{ x } 10^{-19} \text{ J}) = 0.43 \text{ x } 10^{-19}$ J;

(c) For electrons, $\lambda = \frac{h}{mu}$;

$$u = \sqrt{\frac{2E_{\text{kinetic}}}{m}} = \sqrt{\frac{2(0.43 \text{ x } 10^{-19} \text{ J})}{9.109 \text{ x } 10^{-31} \text{ kg}}} = 3 \text{ x } 10^{5} \text{ m/s};$$

$$\lambda = \frac{(6.626 \text{ x } 10^{-34} \text{ kg m}^{2}\text{/s})}{(9.109 \text{ x } 10^{-31} \text{ kg})(3 \text{ x } 10^{5} \text{ m/s})} = 2 \text{ x } 10^{-9} \text{ m or } 2 \text{ nm.}$$

7.32 To determine the wavelength of an electron from the kinetic energy, first determine the speed of the electron and then convert to wavelength. The speed of an electron is determined using $u = \sqrt{\frac{2E_{\text{kinetic}}}{m}}$. Remember, for the units to work out, mass must be in kg and the energy in J such that $\frac{\text{J}}{\text{kg}} = \frac{\text{kg m}^2/\text{s}^2}{\text{kg}} = \text{m}^2/\text{s}^2$. For wavelength calculations use $\lambda = \frac{h}{mu}$:

(a) $E_{\text{electron}} = \left(\frac{76.5\text{ J}}{1\text{ mol}}\right)\left(\frac{1\text{ mol}}{6.022 \times 10^{23}\text{ electrons}}\right) = 1.27 \times 10^{-22}$ J;

$u = \sqrt{\frac{2(1.27 \times 10^{-22}\text{ J})}{9.109 \times 10^{-31}\text{ kg}}} = 1.67 \times 10^4$ m/s;

$\lambda = \frac{6.626 \times 10^{-34}\text{ J s}}{(9.109 \times 10^{-31}\text{ kg})(1.67 \times 10^4\text{ m/s})} = 4.36 \times 10^{-8}$ m or 43.6 nm.

(b) $u = \sqrt{\frac{2(4.77 \times 10^{-18}\text{ J})}{9.109 \times 10^{-31}\text{ kg}}} = 3.24 \times 10^6$ m/s;

$\lambda = \frac{6.626 \times 10^{-34}\text{ J s}}{(9.109 \times 10^{-31}\text{ kg})(3.24 \times 10^6\text{ m/s})} = 2.25 \times 10^{-10}$ m or 0.225 nm.

7.80 Energies and frequencies for transitions in hydrogen atoms can be calculated from the equation for hydrogen atom energy levels:

$$E_n = \frac{-2.18 \times 10^{-18}\text{ J}}{n^2}$$

$$\Delta E_{7-3} = (E_7 - E_3) = \frac{-2.18 \times 10^{-18}\text{ J}}{7^2} - \frac{-2.18 \times 10^{-18}\text{ J}}{3^2} = 1.98 \times 10^{-19}\text{ J}$$

$$\lambda = \frac{hc}{E} = \frac{(6.626 \times 10^{-34}\text{ Js})(2.998 \times 10^8\text{ m/s})}{1.98 \times 10^{-19}\text{ J}} = 1.00 \times 10^{-6}\text{ m}$$

$$\lambda = 1.00 \times 10^{-6}\text{m}\left(\frac{10^6\,\mu\text{m}}{1\text{ m}}\right) = 1.00\ \mu\text{m}$$

These photons lie in the IR, just short of the visible spectral region.

Or use the Rydberg equation,

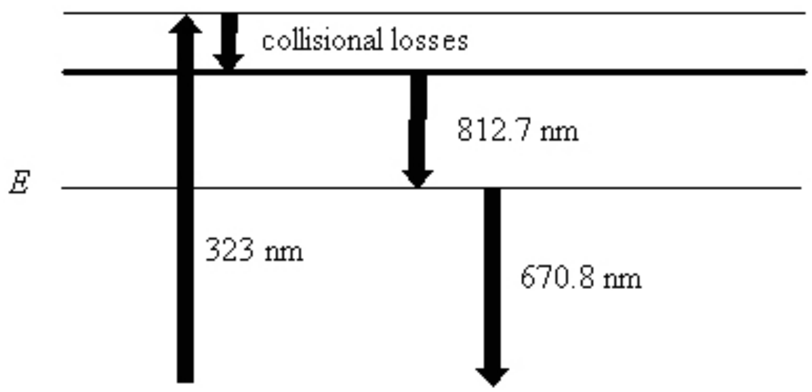
$$\frac{1}{\lambda} = R_H\left(\frac{1}{n_{low}^2} - \frac{1}{n_{high}^2}\right) = 1.0974 \times 10^7\left(\frac{1}{3^2} - \frac{1}{7^2}\right) = 9.9537 \times 10^5\text{ m}^{-1}$$

$\lambda = 1.00 \times 10^{-6}$m

7.86 To determine the energies of the various levels, use the equation $E = \frac{hc}{\lambda}$.

$$E_{\text{photon}} = \frac{hc}{\lambda} = \left(\frac{(6.626 \times 10^{-34}\text{ Js})(2.998 \times 10^{8}\text{ m/s})}{\lambda(\text{in nm})}\right)\left(\frac{10^{9}\text{ nm}}{1\text{m}}\right) = \frac{1.986 \times 10^{-16}\text{ J nm}}{\lambda(\text{in nm})}$$

λ (nm)	323	812.7	670.8
E (10^{-19} J/mol)	6.15	2.444	2.961



$E_{\text{lost}} = [6.15 - (2.444 + 2.961)] \times 10^{-19}\text{ J} = 7.5 \times 10^{-20}\text{ J}$;

Fraction lost $= \frac{7.5 \times 10^{-20} J}{6.15 \times 10^{-19} J} = 1.2 \times 10^{-1}$ (or 12%)

7.38 Remember that $\boldsymbol{n}$ must be a positive integer, $\boldsymbol{l}$ is restricted to zero and positive integers less than $\boldsymbol{n}$, $\boldsymbol{m_l}$ is restricted to integers between $-\boldsymbol{l}$ and $+\boldsymbol{l}$, and $\boldsymbol{m_s} = +1/2$ or $-1/2$. For $\boldsymbol{m_l} = -2$, $\boldsymbol{l}$ must be a positive integer ≥ 2, requiring that $\boldsymbol{n}$ be a positive integer $\geq \boldsymbol{l} + 1$.

7.40 Remember that $\boldsymbol{n}$ must be a positive integer, $\boldsymbol{l}$ is restricted to zero and positive integers less than $\boldsymbol{n}$, $\boldsymbol{m_l}$ is restricted to integers between $-\boldsymbol{l}$ and $+\boldsymbol{l}$, and $\boldsymbol{m_s} = +1/2$ or $-1/2$.
(a) non-existent: $\boldsymbol{l}$ must be positive; (b) actual; (c) non-existent: $\boldsymbol{m_l}$ cannot be larger than $\boldsymbol{l}$; and (d) actual.

7.46 The d_{xz} looking down the...