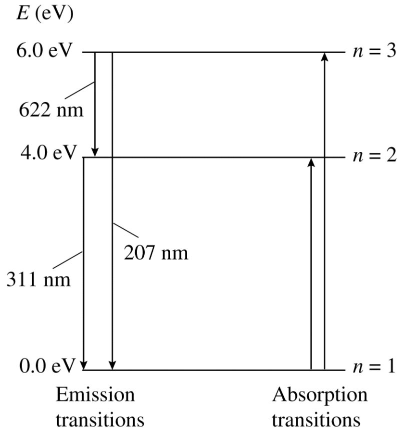
38.24. Model: To conserve energy, the emission and the absorption spectra must have exactly the energy lost or gained by the atom in the appropriate quantum jumps.

Solve: (a)



(b) From Equation 38.4, the energy of a light quantum is $E = hf = hc/\lambda$. We can use this equation to find the emission and absorption wavelengths. The emission energies from the above energy-level diagram are: $E_{2\to1} = 4.0$ eV, $E_{3\to1} = 6.0$ eV, and $E_{3\to2} = 2.0$ eV. The wavelength corresponding to the 2→1 transition is

$$\lambda_{2\to1} = \frac{hc}{E_{21}} = \frac{(4.14\times10^{-15}\ \text{eV s})(3.0\times10^{8}\ \text{m/s})}{4.0\ \text{eV}} = 311\ \text{nm}$$

Likewise, $\lambda_{3\to1} = hc/E_{3\to1} = 207\ \text{nm}$, and $\lambda_{3\to2} = 622$ nm.

(c) Absorption transitions start from the $n = 1$ ground state. The energies in the atom's absorption spectrum are $E_{1\to2} = 4.0$ eV and $E_{1\to3} = 6.0$ eV. The corresponding wavelengths are $\lambda_{1\to2} = hc/E_{1\to2} = 311$ nm and $\lambda_{1\to3} = hc/E_{1\to3} = 207$ nm.