

7 • Atomic Structure

CALCULATION PRACTICE — 2

Formulas and Constants

$c = \lambda \nu$	$E = h\nu$	$E = \frac{hc}{\lambda}$	$E_n = -\frac{Rhc}{n^2}$	$\lambda = \frac{h}{mv}$	$\frac{1}{\lambda} = R \left(\frac{1}{2^2} - \frac{1}{n^2} \right)$
$c = 3.00 \times 10^8 \text{ m/s}$	$h = 6.626 \times 10^{-34} \text{ J}\cdot\text{s}$	$Rhc = 2.18 \times 10^{-18} \text{ J}$	$R = 1.0974 \times 10^7 \text{ m}^{-1}$		

1. Sketch the electron energy levels ($n=1$ through $n=5$) for the hydrogen atom.



2. Calculate the energy of an electron in the $n=2$ energy level of hydrogen. Calculate the energy of an electron in the $n=3$ energy level. What is the difference in energy of these two levels? If a photon of light had this energy, what would its wavelength be?

$$E_2 = -\frac{2.18 \times 10^{-18} \text{ J}}{4} = -5.45 \times 10^{-19} \text{ J}$$

$$\Delta = 3.03 \times 10^{-19} \text{ J}$$

$$E_3 = -\frac{2.18 \times 10^{-18} \text{ J}}{9} = -2.42 \times 10^{-19} \text{ J}$$

$$\lambda = \frac{hc}{E} = \frac{(6.626 \times 10^{-34} \text{ J}\cdot\text{s})(2.998 \times 10^8 \text{ m/s})}{3.03 \times 10^{-19} \text{ J}} = 6.56 \times 10^{-7} \text{ m}$$

3. Use the Rydberg equation above to calculate the wavelength of a photon when $n=3$. How does this compare with your answer in question 2?

$$\frac{1}{\lambda} = 1.0974 \times 10^7 \text{ m}^{-1} \left(\frac{1}{2^2} - \frac{1}{3^2} \right)$$

$$\lambda = 6.56 \times 10^{-7} \text{ m}$$

Same

4. An electron moves from the $n=5$ to the $n=1$ quantum level and emits a photon with an energy of $2.093 \times 10^{-18} \text{ J}$. How much energy must the atom absorb to move an electron from $n=1$ to $n=5$? What is the wavelength of this energy?

$$\text{the same amount}$$

$$\lambda = \frac{hc}{E} = \frac{(6.626 \times 10^{-34} \text{ J}\cdot\text{s})(2.998 \times 10^8 \text{ m/s})}{2.093 \times 10^{-18} \text{ J}} = 9.5 \times 10^{-8} \text{ m} = 9.49 \times 10^{-8} \text{ m}$$

5. An electron moves with a velocity of $2.5 \times 10^8 \text{ cm/s}$. What is its wavelength? (The mass of an electron is $9.109 \times 10^{-28} \text{ g}$.)

$$\lambda = \frac{h}{mv} = \frac{6.626 \times 10^{-34} \text{ J}\cdot\text{s}}{(9.109 \times 10^{-28} \text{ g} \times \frac{1 \text{ kg}}{10^3 \text{ g}})(2.5 \times 10^8 \frac{\text{cm}}{\text{s}} \times \frac{1 \text{ m}}{100 \text{ cm}})} = 2.9 \times 10^{-10} \text{ m}$$

6. Calculate the wavelength (in nanometers) associated with a $1.0 \times 10^2 \text{ g}$ golf ball moving at $30. \text{ m/s}$ (about 67 mph). How fast must the ball travel to have a wavelength of $5.6 \times 10^{-3} \text{ nm}$?

$$\lambda = \frac{h}{mv} = \frac{6.626 \times 10^{-34} \text{ J}\cdot\text{s}}{(1.0 \times 10^2 \text{ g})(\frac{1 \text{ kg}}{10^3 \text{ g}})(30 \frac{\text{m}}{\text{s}})} = 2.2 \times 10^{-34} \text{ m} = 2.2 \times 10^{-25} \text{ nm}$$

$$5.6 \times 10^{-3} \text{ nm} = \frac{6.626 \times 10^{-34} \text{ J}\cdot\text{s}}{(1.0 \text{ kg}) \times v}$$

$$v = \frac{6.626 \times 10^{-34} \text{ J}\cdot\text{s}}{(1.0 \text{ kg})(5.6 \times 10^{-3} \text{ nm})(\frac{1 \text{ m}}{10^9 \text{ nm}})} = 1.18 \times 10^{-21} \frac{\text{m}}{\text{s}}$$

$$= 1.18 \times 10^{-21} \frac{\text{m}}{\text{s}}$$

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