

# 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?

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

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

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

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

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?

The same amount  
 $-2.093 \times 10^{-18} \text{ J}$

$$\lambda = \frac{hc}{E} = \frac{(6.626 \times 10^{-34} \text{ J}\cdot\text{s})(2.998 \times 10^8 \text{ m/s})}{2.093 \times 10^{-18} \text{ J}} = 9.5 \times 10^{-8} \text{ m} = 9.491 \times 10^{-8} \text{ m}$$

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  $67 \text{ mph}$ ). How fast must the ball travel to have a wavelength of  $5.6 \times 10^{-3} \text{ nm}$ ?

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

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

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

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

# 7 Atomic Structure Calculation Practice 1 Answers

**JG Myers**



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