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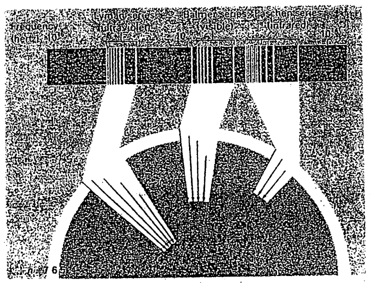
5 INTERPRETING GRAPHICS

Use with Section 5.3 → p 139, Figure 5.10



* Formulas

$E = h \cdot \nu$

↳ $6.626 \cdot 10^{-34}$ J·s

$c = \lambda \nu$

↳ 3.0×10^{8} m/sec

Figure 1 The emission spectrum and orbit-transition diagram for hydrogen.

Table 1

Transition	E (J)	ν (s^{-1})	λ(m)	Type of Radiation
$n = 6 \rightarrow n = 5$	2.66×10^{-20}	4.01×10^{13}	7.48×10^{-6}	IR
$n = 6 \rightarrow n = 4$	7.57×10^{-20}	1.14×10^{14}	2.63×10^{-6}	IR
$n = 6 \rightarrow n = 3$	1.82×10^{-19}	2.75×10^{14}	1.09×10^{-6}	IR
$n = 6 \rightarrow n = 2$	4.84×10^{-19}	7.30×10^{14}	4.11×10^{-7}	Visible - Violet
$n = 6 \rightarrow n = 1$	2.12×10^{-18}	3.20×10^{15}	9.40×10^{-8}	UV
$n = 5 \rightarrow n = 4$	4.91×10^{-20}	7.41×10^{13}	4.05×10^{-6}	IR
$n = 5 \rightarrow n = 3$	1.55×10^{-19}	2.34×10^{14}	1.28×10^{-6}	IR
$n = 5 \rightarrow n = 2$	4.56×10^{-19}	6.88×10^{14}	4.36×10^{-7}	Visible - Indigo
$n = 5 \rightarrow n = 1$	2.09×10^{-18}	3.15×10^{15}	9.50×10^{-8}	UV
$n = 4 \rightarrow n = 3$	1.06×10^{-19}	1.60×10^{14}	1.88×10^{-6}	IR
$n = 4 \rightarrow n = 2$	4.09×10^{-19}	6.17×10^{14}	4.86×10^{-7}	Visible - Blue
$n = 4 \rightarrow n = 1$	2.04×10^{-18}	3.08×10^{15}	6.20×10^{-8}	UV
$n = 3 \rightarrow n = 2$	3.03×10^{-19}	4.57×10^{14}	6.56×10^{-7}	Visible - Orange
$n = 3 \rightarrow n = 1$	1.94×10^{-18}	2.93×10^{15}	1.02×10^{-7}	Visible / UV
$n = 2 \rightarrow n = 1$	1.64×10^{-18}	2.48×10^{15}	1.21×10^{-7}	Visible / UV

Red
700 nm
7.00×10^{-7} m
(0.700×10^{-6})

Violet
400 nm
4.00×10^{-7} m
(0.400×10^{-6})

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1. Figure 1 summarizes the quantum model of the hydrogen atom originally proposed by Neils Bohr to account for the interaction of hydrogen with electromagnetic radiation. The energy changes associated with each electron transition for the lowest six energy levels of hydrogen are listed in Table 1. Calculate the frequency of the emitted radiation for each transition.

2. Calculate the wavelength in meters for each energy level transition and fill in the column for wavelength.

3. Determine the type of radiation (ultraviolet, visible, or infrared) that corresponds to each wavelength.

4. Which transitions result in the emission of visible light?

   Blue/Violet $n = 6 \rightarrow n = 2$ (?) $n = 3 \rightarrow n = 1$
   $n = 5 \rightarrow n = 2$ (?) $n = 2 \rightarrow n = 1$
   $n = 4 \rightarrow n = 2$
   $n = 3 \rightarrow n = 2$

5. If the wavelengths of blue, green, and red light are approximately 400 nm, 500 nm, and 650 nm, respectively, what colors in the visible spectrum correspond to the transitions stated in your answer to question 4?

   blue: 400nm
   green: 500 nm
   red: 650 nm

6. What is the common feature among transitions where the resulting radiation lies within the visible light range of the electromagnetic spectrum?

   • All have λ of 10^{-7} m

7. The Bohr model, although historically important, was limited in its ability to explain the behavior of more complex elements and ions. To which of the following atoms or ions would you expect the Bohr model to apply?

   Be, He^{+}, K, Li^{2+}

   Bohr studied H w/ 1e⁻ → He^{+} has 1e⁻