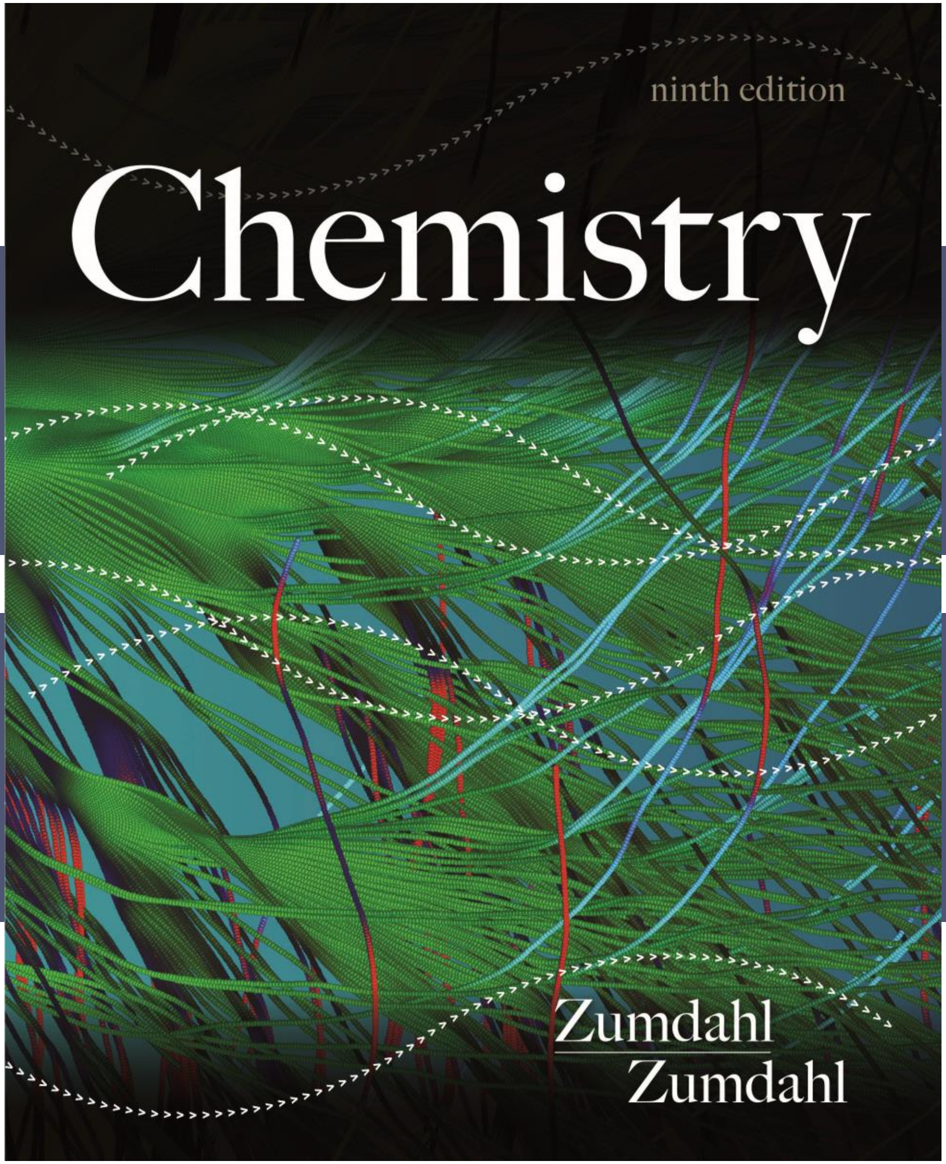



# Chapter 7

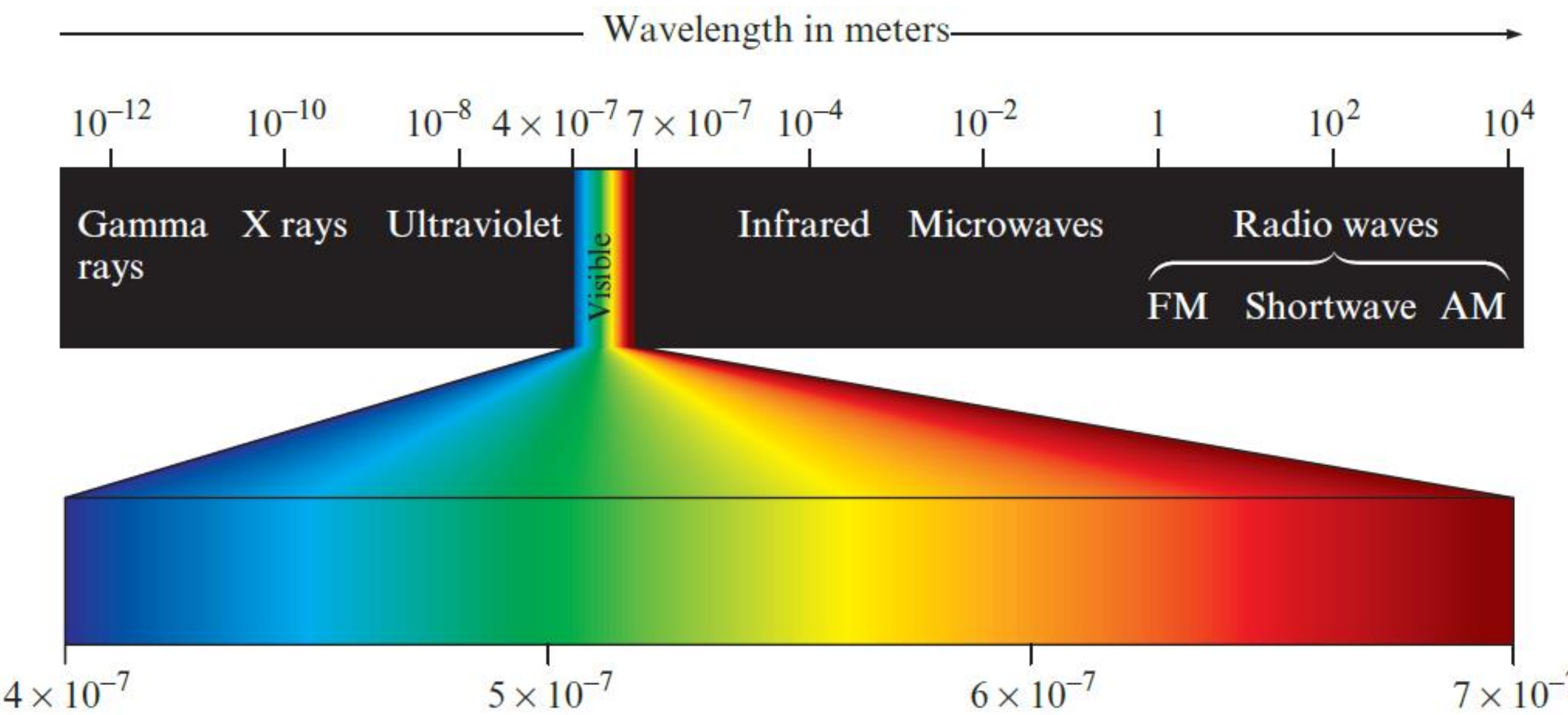
## *Atomic Structure and Periodicity*

# Relationship between Wavelength and Frequency

- $\lambda$ - Wavelength in meters
- $\nu$ - Frequency in cycles per second
- $c$ - Speed of light ($2.9979 \times 10^8$ m/s)

$$\lambda \nu = c$$

## **Figure 7.2** - Classification of Electromagnetic Radiation



## Write question and answer in your notes, compare with partner

- The brilliant red colors seen in fireworks are due to the emission of light with wavelengths around 650 nm when strontium salts are heated
  - Calculate the frequency of red light of wavelength $6.50 \times 10^2$ nm

## Interactive Example 7.1 - Solution (Continued)

- Changing the wavelength to meters, we have

$$6.50 \times 10^2 \ \cancel{\text{nm}} \times \frac{1 \text{ m}}{10^9 \ \cancel{\text{nm}}} = 6.50 \times 10^{-7} \text{ m}$$

- And

$$\nu = \frac{c}{\lambda} = \frac{2.9979 \times 10^8 \ \cancel{\text{m}}/\text{s}}{6.50 \times 10^{-7} \ \cancel{\text{m}}}$$

$$= 4.61 \times 10^{14} \text{ s}^{-1}$$

$$= 4.61 \times 10^{14} \text{ Hz}$$

## Max Planck

- Change in energy ($\Delta E$) can be represented as follows:

$$\Delta E = nh\nu$$

  - $n$ - Integer
  - $h$ - Planck's constant
  - $\nu$ - Frequency of electromagnetic radiation absorbed or emitted

## Write question and answer in your notes, compare with partner

- The blue color in fireworks is often achieved by heating copper(I) chloride (CuCl) to about 1200° C
  - Then the compound emits blue light having a wavelength of 450 nm
  - What is the increment of energy (the quantum) that is emitted at $4.50 \times 10^2$ nm by CuCl?

# Interactive Example 7.2 - Solution

- The quantum of energy can be calculated from the following equation:

$$\Delta E = h\nu$$

  - The frequency $\nu$ for this case can be calculated as follows:

$$\nu = \frac{c}{\lambda} = \frac{2.9979 \times 10^{8}\ \cancel{\text{m}}/\text{s}}{4.50 \times 10^{-7}\ \cancel{\text{m}}} = 6.66 \times 10^{14}\ \text{s}^{-1}$$

## Interactive Example 7.2 - Solution (Continued)

- Therefore,

$$\Delta E = h\nu = \left(6.626 \times 10^{-34}\ \text{J} \cdot \cancel{\text{s}}\right)\left(6.66 \times 10^{14}\ \cancel{\text{s}^{-1}}\right)$$

$$= 4.41 \times 10^{-19}\ \text{J}$$

- A sample of CuCl emitting light at 450 nm can lose energy only in increments of $4.41 \times 10^{-19}$ J, the size of the quantum in this case

# Albert Einstein

- $h$ - Planck's constant
- $\nu$ - Frequency of radiation
- λ - Wavelength of radiation

$$E_{\text{photon}} = h\nu = \frac{hc}{\lambda}$$

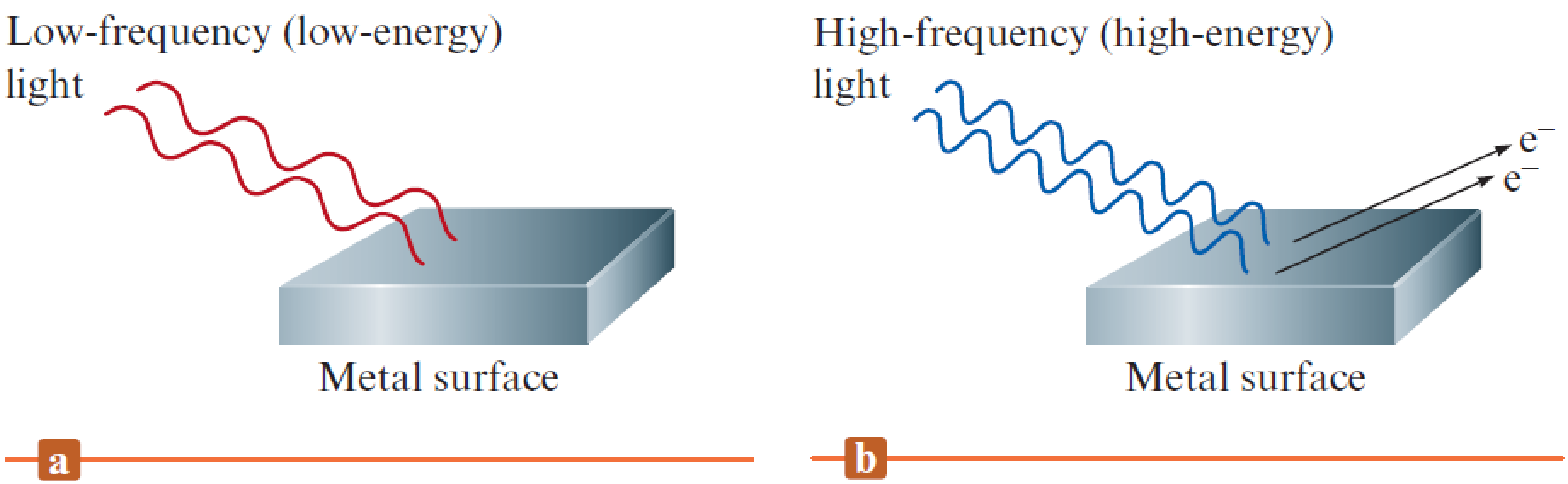
## Photoelectric Effect

- Electrons are emitted from surface of metal when light strikes it
- Observations
  - When varying $v$ , no e- emitted by given metal below threshold frequency ($v_0$)
  - When $v < v_0$, no e- emitted, regardless of the intensity

## Photoelectric Effect (Continued 1)

- When $v > v_0$:
  - The number of e- emitted increases with the intensity
  - The kinetic energy (KE) of the emitted e- increases linearly with the frequency of the light
- Assumptions
  - Electromagnetic radiation is quantized
  - $v_0$ represents the minimum E required to remove e- from surface of metal

## Figure 7.4 - The Photoelectric Effect



## Photoelectric Effect (Continued 2)

- When $v > v_0$, excess energy after removal of the e- is given to the e- as kinetic energy (KE)

$$\mathrm{KE}_{\text{electron}} = \frac{1}{2} m \upsilon^2 = h\nu - h\nu_0$$

  - $m$ - Mass of electron
  - $\upsilon^2$ - Velocity of e-
  - $h\nu$ - E of incident photon
  - $h\nu_0$ – E- to remove e- from metal’s surface

## Einstein's Theory of Relativity

- Einstein proposed that energy has mass

$$E = mc^2$$

## Louis de Broglie

- Ascertained if matter that is assumed to be particulate exhibits wave properties

$$m = \frac{h}{\lambda \upsilon} \longleftarrow$$

Relationship between mass and wavelength for electromagnetic radiation

  - used to calculate the wavelength of a particle

## Write and answer the question in your notes, compare with partner

- Compare the wavelength for an electron (mass = $9.11 \times 10^{-31}$ kg) traveling at a speed of $1.0 \times 10^{7}$ m/s with that for a ball (mass = 0.10 kg) traveling at 35 m/s

## Interactive Example 7.3 - Solution

- For the electron,

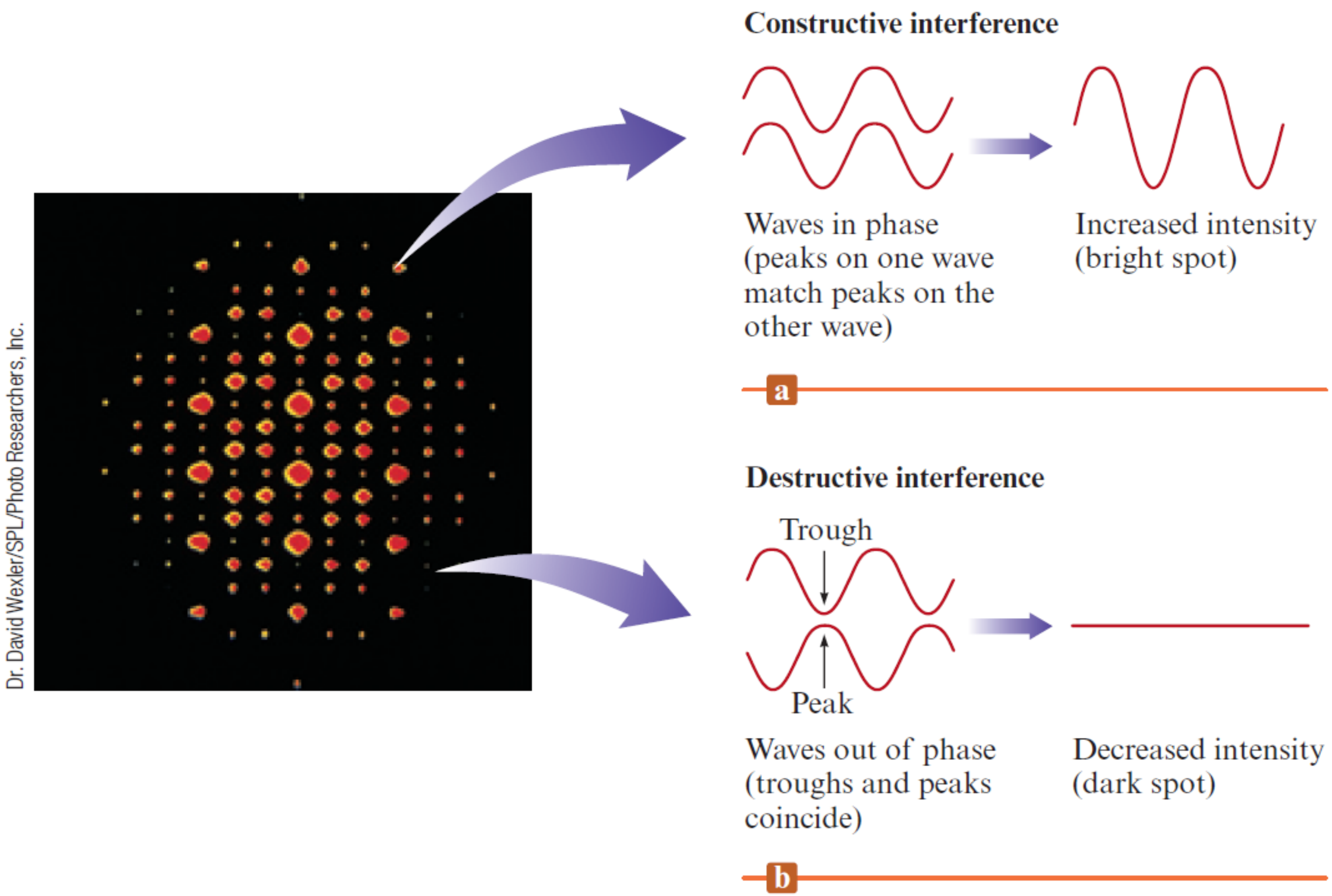
$$\lambda_e = \frac{6.626 \times 10^{-34}\ \dfrac{\cancel{\text{kg}} \cdot \cancel{\text{m}} \cdot \text{m}}{\cancel{\text{s}}}}{\left(9.11 \times 10^{-31}\ \cancel{\text{kg}}\right)\left(1.0 \times 10^{7}\ \cancel{\text{m}}/\cancel{\text{s}}\right)} = 7.27 \times 10^{-11}\ \text{m}$$

## Interactive Example 7.3 - Solution (Continued)

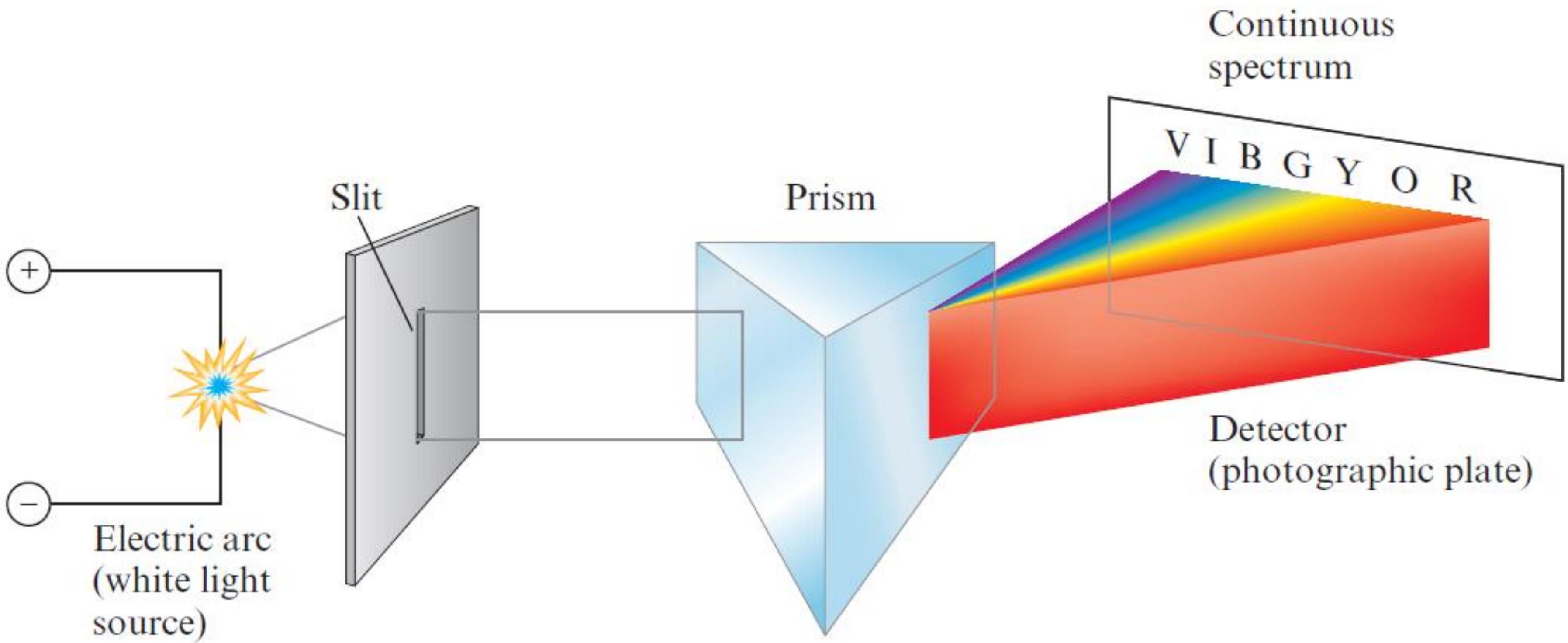
- For the ball,

$$\lambda_b = \frac{6.626 \times 10^{-34} \dfrac{\cancel{\text{kg}} \cdot \cancel{\text{m}} \cdot \text{m}}{\cancel{\text{s}}}}{\left(0.10\ \cancel{\text{kg}}\right)\left(35\ \cancel{\text{m}}/\cancel{\text{s}}\right)} = 1.9 \times 10^{-34}\ \text{m}$$

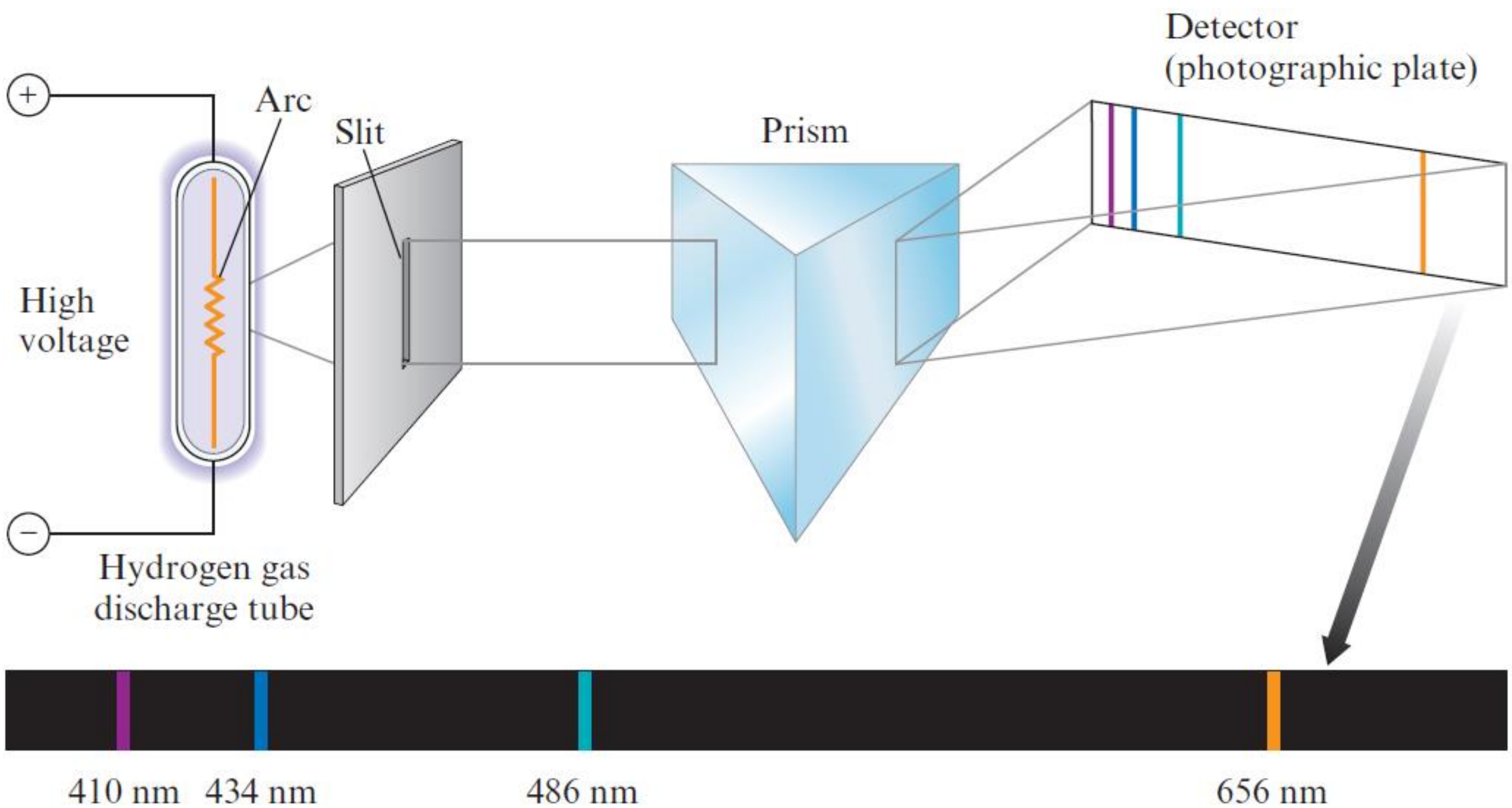
# Figure 7.6 - Diffraction Pattern of a Beryl Crystal



## Figure 7.7 (a) - A Continuous Spectrum

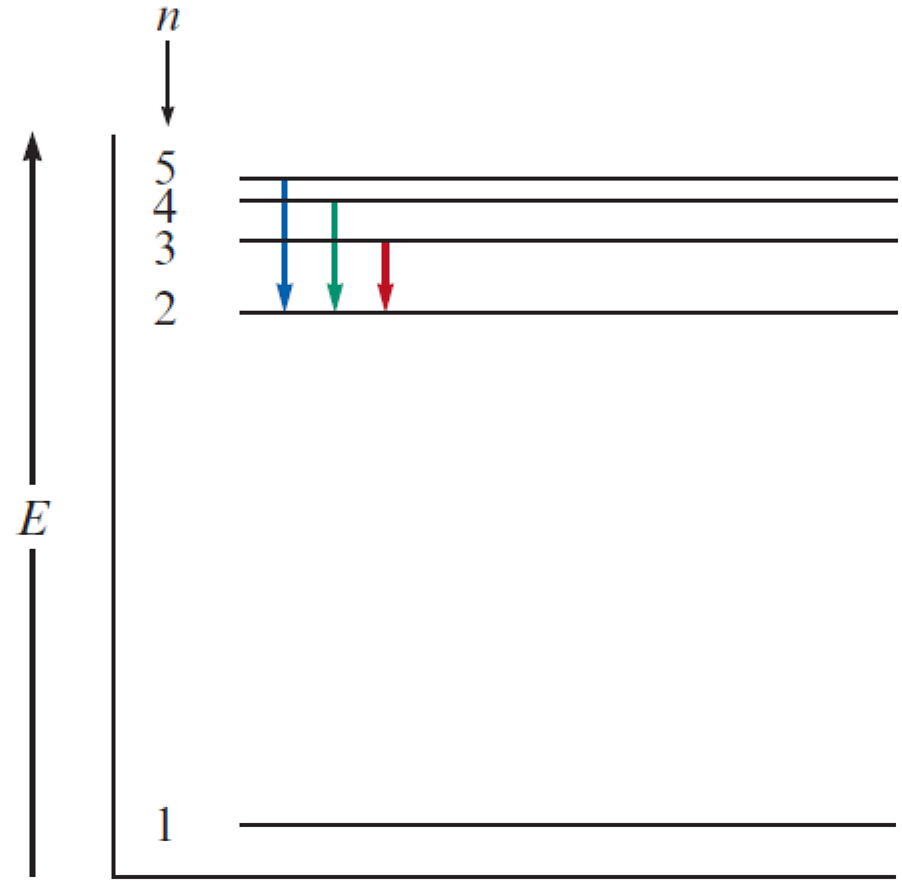


# Figure 7.7 (b) - The Hydrogen Line Spectrum



## **Figure 7.9 (a)** - An Energy-Level Diagram for Electronic Transitions

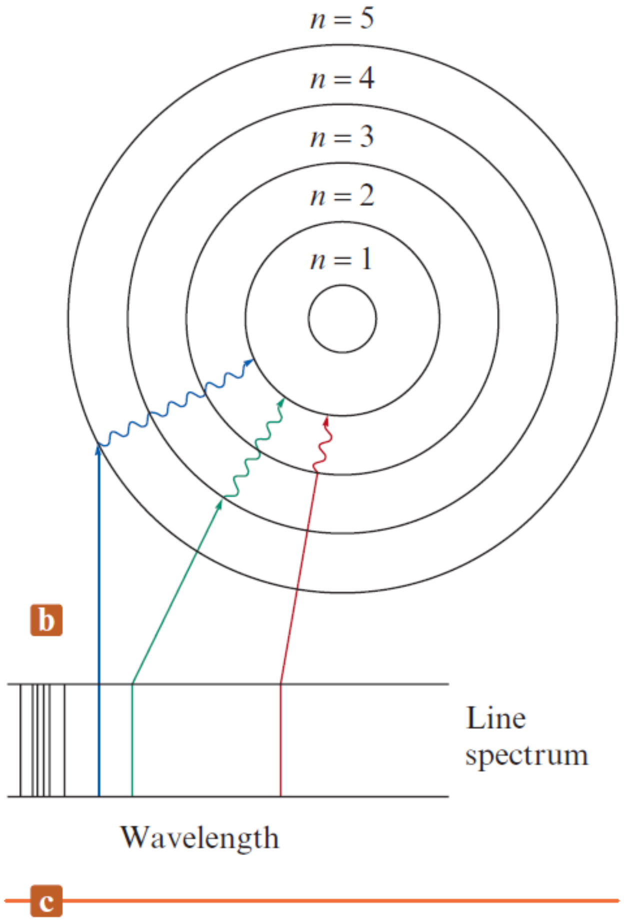
- Bohr's model gave hydrogen atom energy levels consistent with the hydrogen emission spectrum



## Figure 7.9 (b and c) - Electronic Transitions in the Bohr Model for the Hydrogen Atom

b) An orbit-transition diagram, which accounts for the experimental spectrum

c) The resulting line spectrum on a photographic plate is shown



# Bohr's Model

- Expression for energy levels available to the electrons in the hydrogen atom

$$E = -2.178 \times 10^{-18}\,\text{J}\left(\frac{Z^2}{n^2}\right)$$

  - *n* - An integer (A large *n* value implies a large orbit radius)
  - *Z* - Nuclear charge

## Calculation of Change in Energy (Δ*E*) and Wavelength of the Emitted Photon

- Calculation of the wavelength of the emitted photon

$$\Delta E = h\left(\frac{c}{\lambda}\right) \text{ or } \lambda = \frac{hc}{\Delta E}$$

## Write and answer the question in your notes, compare with partner

- Calculate the energy required to excite the hydrogen electron from level $n = 1$ to level $n = 2$
  - Also calculate the wavelength of light that must be absorbed by a hydrogen atom in its ground state to reach this excited state

# Interactive Example 7.4 - Solution

- Use the following equation, with $Z = 1$:

$$E = -2.178 \times 10^{-18}\,\text{J}\left(\frac{Z^2}{n^2}\right)$$

$$E_1 = -2.178 \times 10^{-18}\,\text{J}\left(\frac{1^2}{1^2}\right) = -2.178 \times 10^{-18}\,\text{J}$$

$$E_2 = -2.178 \times 10^{-18}\,\text{J}\left(\frac{1^2}{2^2}\right) = -5.445 \times 10^{-19}\,\text{J}$$

# Interactive Example 7.4 - Solution (Continued 1)

$$\Delta E = E_2 - E_1 = \left(-5.445 \times 10^{-19}\text{J}\right) - \left(-2.178 \times 10^{-18}\text{J}\right)$$
$$= 1.633 \times 10^{-18}\ \text{J}$$

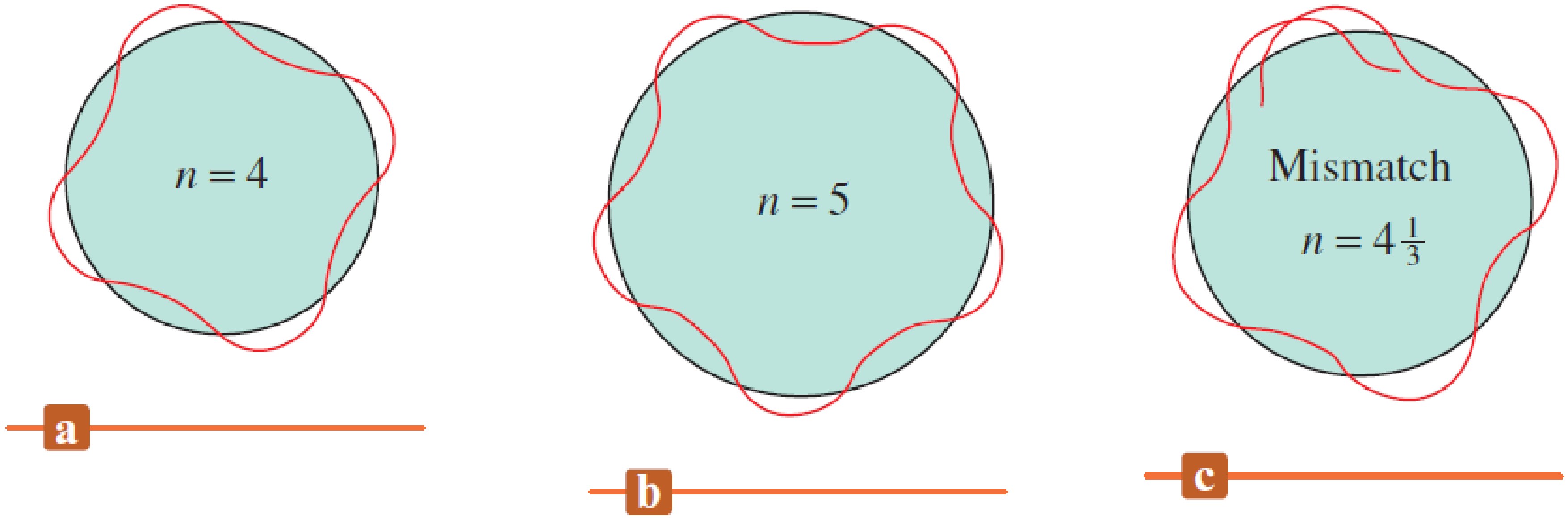
- The positive value for Δ*E* indicates that the system has gained energy
  - The wavelength of light that must be absorbed to produce this change can be calculated using $\lambda = hc/\Delta E$

# Interactive Example 7.4 - Solution (Continued 2)

$$\lambda = \frac{hc}{\Delta E} = \frac{\left(6.626 \times 10^{-34}\ \cancel{J} \cdot \cancel{s}\right)\left(2.9979 \times 10^{8}\ \text{m}/\cancel{s}\right)}{1.633 \times 10^{-18}\ \cancel{J}}$$

$$\lambda = 1.216 \times 10^{-7}\ \text{m}$$

# Figure 7.11 - Hydrogen Electron Visualized as a Standing Wave



## Erwin Schrödinger and Quantum Mechanics

- Schrödinger's equation

$$\hat{H}\psi = E\psi$$

  - $\psi$ - **Wave function**
    - Function of the coordinates of the electron's position in three-dimensional space
  - $\hat{H}$ - Operator
    - Contains mathematical terms that produce the total energy of an atom when applied to the wave function

## Heisenberg's Uncertainty Principle

- There is a fundamental limitation to just how precisely we can know both the position and momentum of a particle at a given time

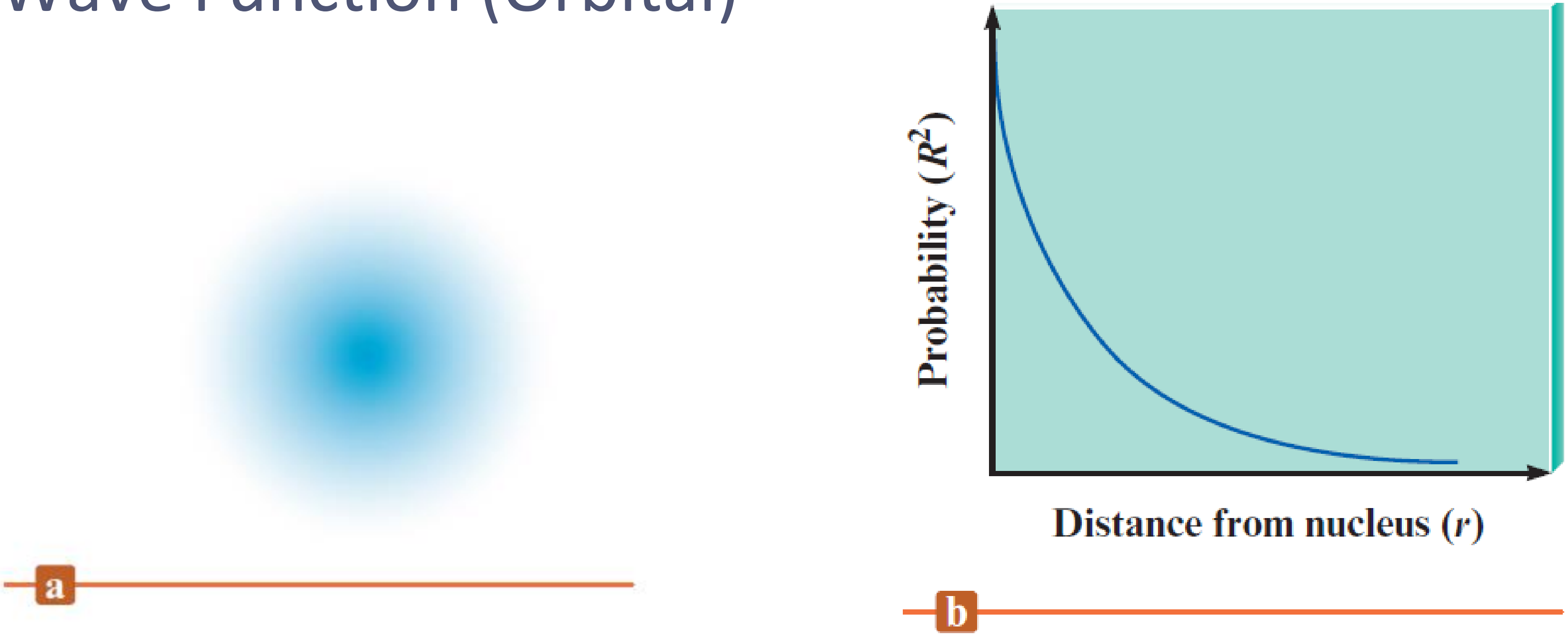
$$\Delta x \cdot \Delta(m\upsilon) \geq \frac{h}{4\pi}$$

  - $\Delta x$ - Uncertainty in a particle's position
  - $\Delta(m\upsilon)$ - Uncertainty in particle momentum
    - Minimum uncertainty in the product $\Delta x \cdot \Delta(m\upsilon)$ is $h/4\pi$
  - $h$ - Planck's constant

## Square of a Wave Function

- Indicates the probability of finding an electron near a particular point in space
- Represented by probability distribution
  - **Probability distribution**: Intensity of color is used to indicate the probability value near a given point in space

# Figure 7.12 - Probability Distribution for the Hydrogen 1*s* Wave Function (Orbital)



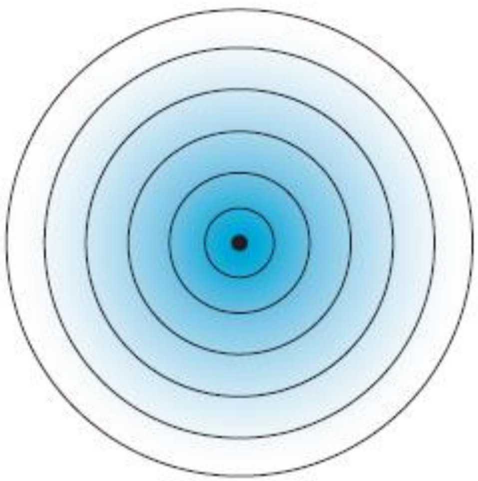

The probability distribution for the hydrogen 1s orbital in three-dimensional space

The probability of finding the electron at points along a line drawn from the nucleus outward in any direction for the hydrogen 1*s* orbital

## Radial Probability Distribution

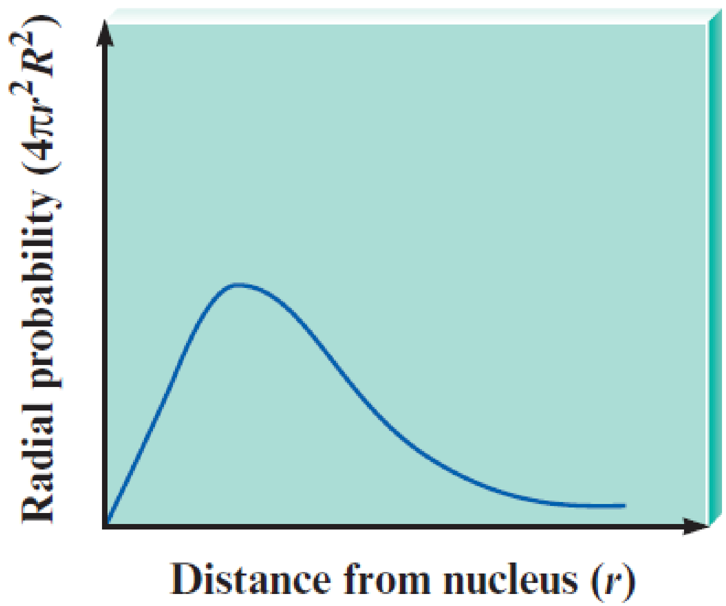
- Plots the total probability of finding an electron in each spherical shell versus the distance from the nucleus
  - Probability of finding an electron at a particular position is greatest near the nucleus
  - Volume of the spherical shell increases with distance from the nucleus

## Figure 7.13 - Radial Probability Distribution



a

Cross section of the hydrogen 1*s* orbital probability distribution divided into successive thin spherical shells




b

Plot of the total probability of finding the electron in each thin spherical shell as a function of distance from the nucleus

## Quantum Numbers

- Series of numbers that express various properties of an orbital
  - Principal quantum number ($n$)
  - Angular momentum quantum number ($\mathrm{l}$)
  - Magnetic quantum number ($m_{\mathrm{l}}$)

## Discuss with your partner, then discuss with another group. With that group, come up with an answer.

- You have learned that each orbital is allowed two electrons, and this pattern is evident on the periodic table
  - What if each orbital was allowed three electrons?
    - How would this change the appearance of the periodic table?

## Answer in your notes, compare with partner

- Give the electron configurations for sulfur (S), cadmium (Cd), hafnium (Hf), and radium (Ra) using the periodic table inside the front cover of this book

# Periodic Trends

- Ionization energy
- Electron affinity
- Atomic radius

## Write and answer the question in your notes, compare with partner

- The first ionization energy for phosphorus is 1060 kJ/mol, and that for sulfur is 1005 kJ/mol
  - Why?

## Example 7.8 - Solution (Continued)

- Ordinarily, the first IE increases across a period, so we expect S to have a greater ionization energy
- However, in this case the fourth *p* electron in S must be placed in an already occupied orbital
    - Repulsions that result cause e- to be more easily removed

## Answer in your notes, compare with partner

- Predict the trend in radius for the following ions:
  - $Be^{2+}$
  - $Mg^{2+}$
  - $Ca^{2+}$
  - $Sr^{2+}$