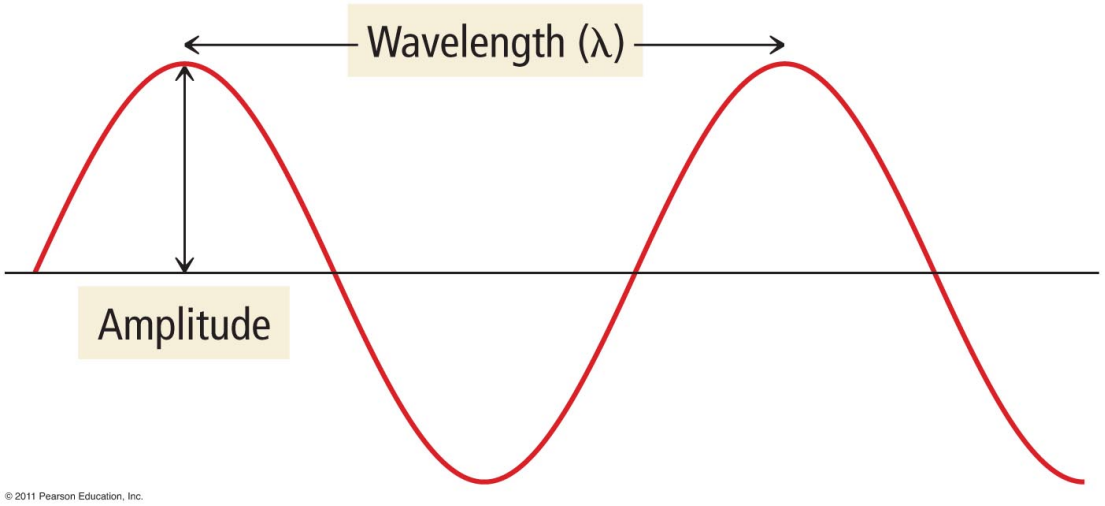
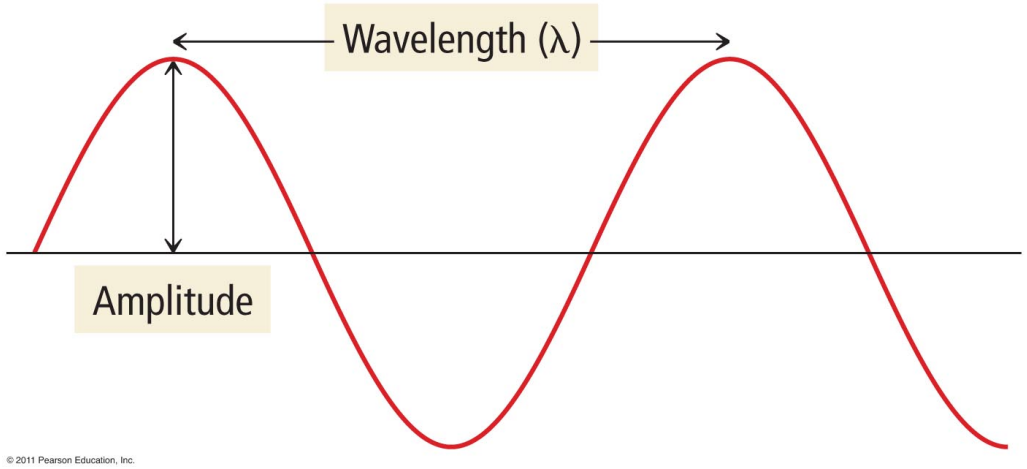
Chapter 7: The Quantum-Mechanical Model of the Atom

7.2 Properties of Waves:



- **Wavelength (λ)**



- **Speed (*u*)**
  - For <u>all</u> electromagnetic radiation:

- **Frequency (ν)**
  - Number of cycles of the wave per second

- **Amplitude**: height of peaks – intensity of wave.

Figure 7.2 Wavelength and Amplitude

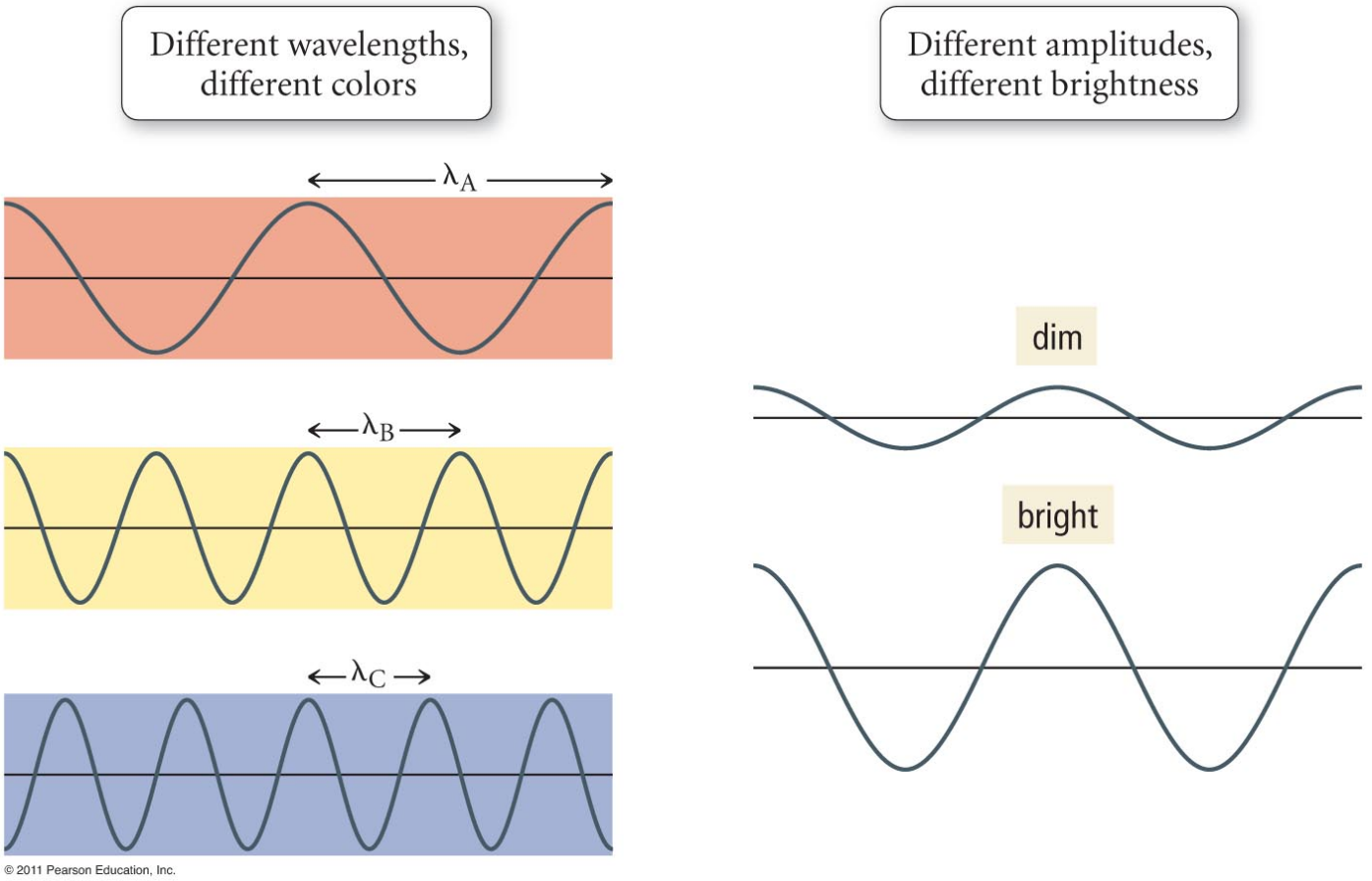
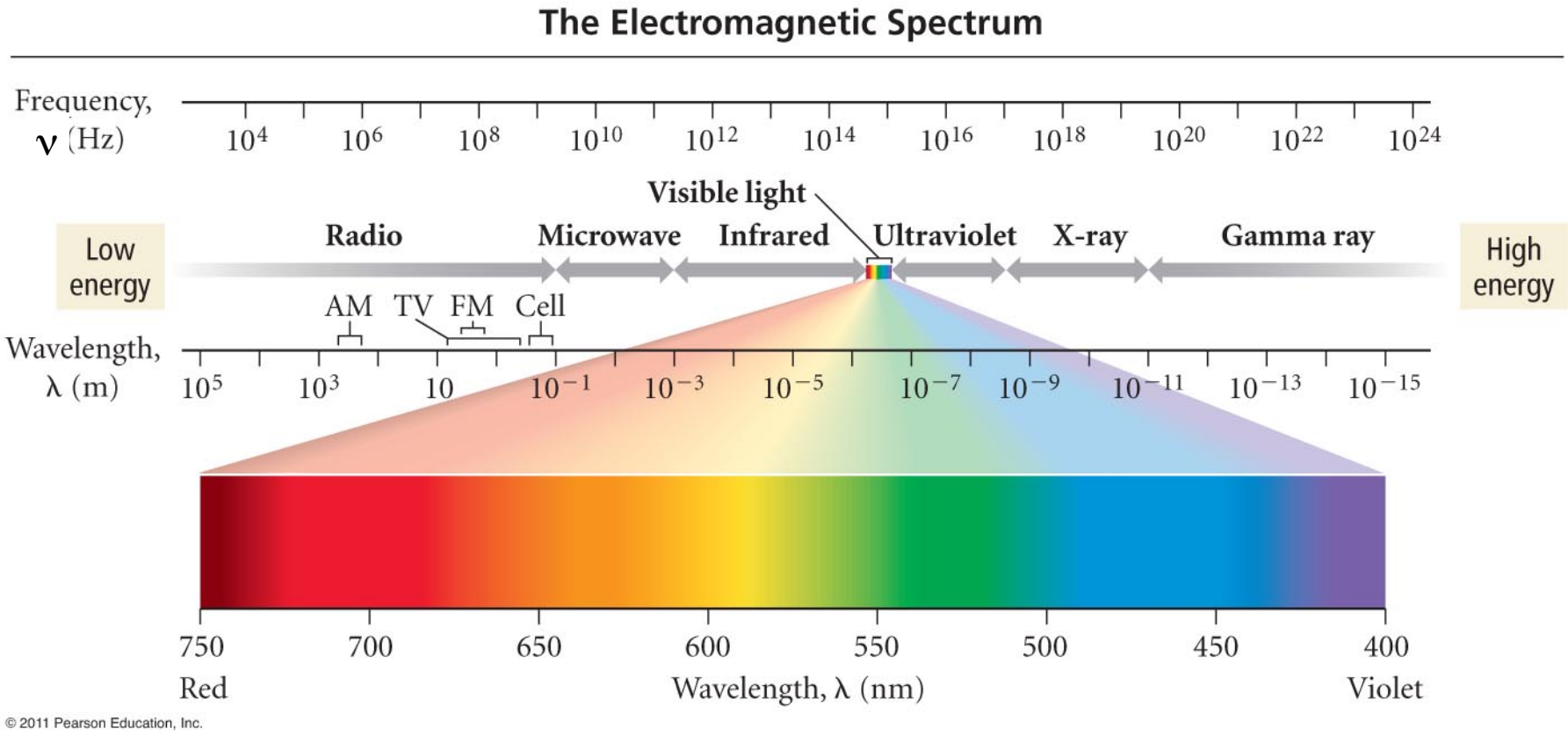


Figure 7.5 The Electromagnetic Spectrum



Quantum Theory

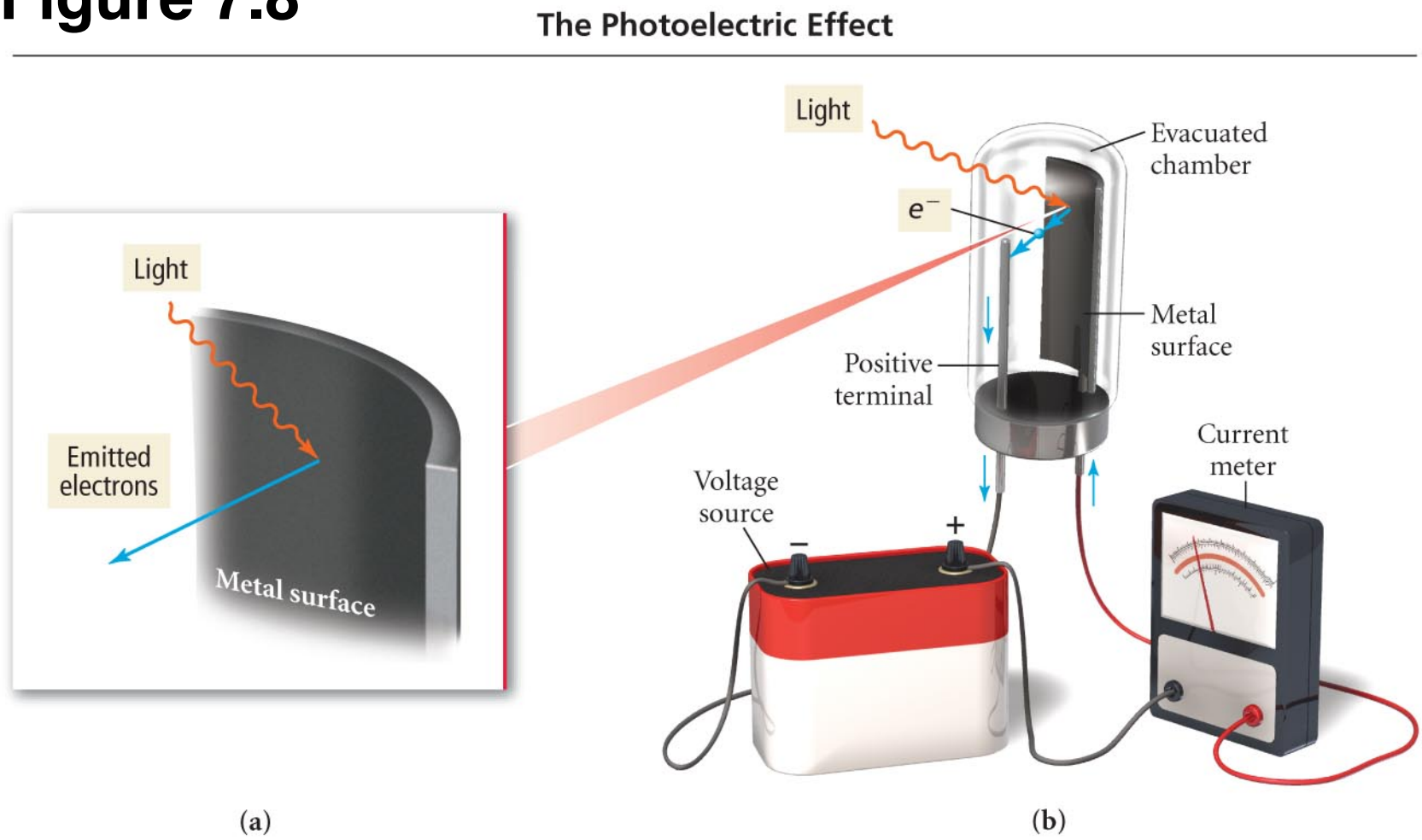
Three phenomena could not be explained by the laws of classical physics:

1) Blackbody radiation (glow of solids when hot)
2) The photoelectric effect
3) Atomic spectra

- Max Planck: a substance can only emit certain amounts (quanta) of energy.
- The amount of energy emitted is proportional to the **frequency** of the light:

h is Planck's constant = 6.626×10^{-34} J•s

Figure 7.8



- Photoelectric effect: Shining light on a metal surface can cause electrons to be ejected from the metal.

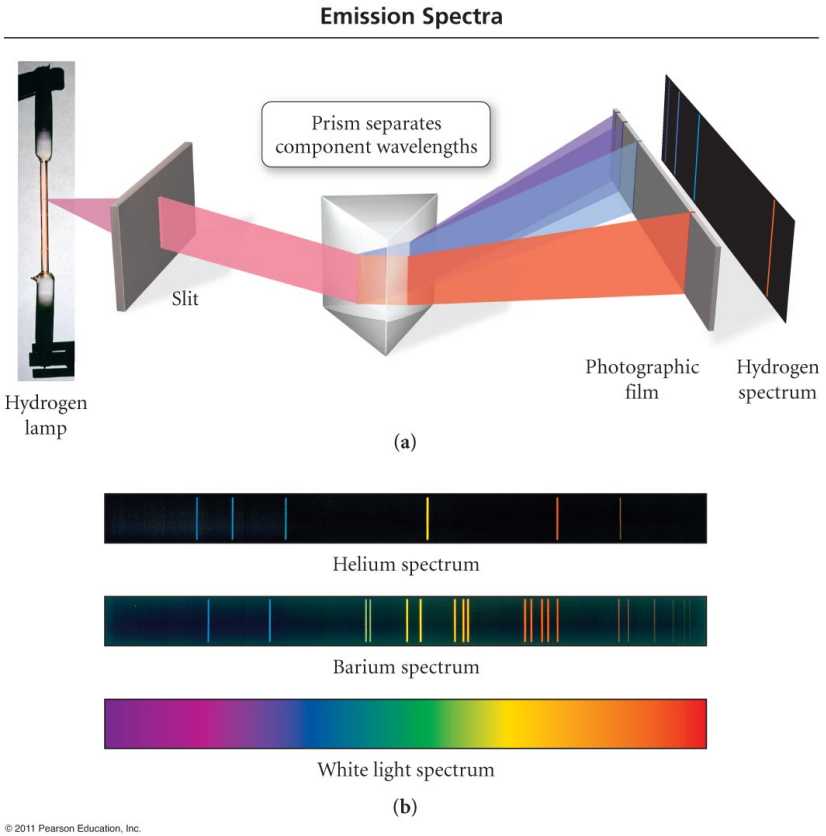
The Photoelectric Effect

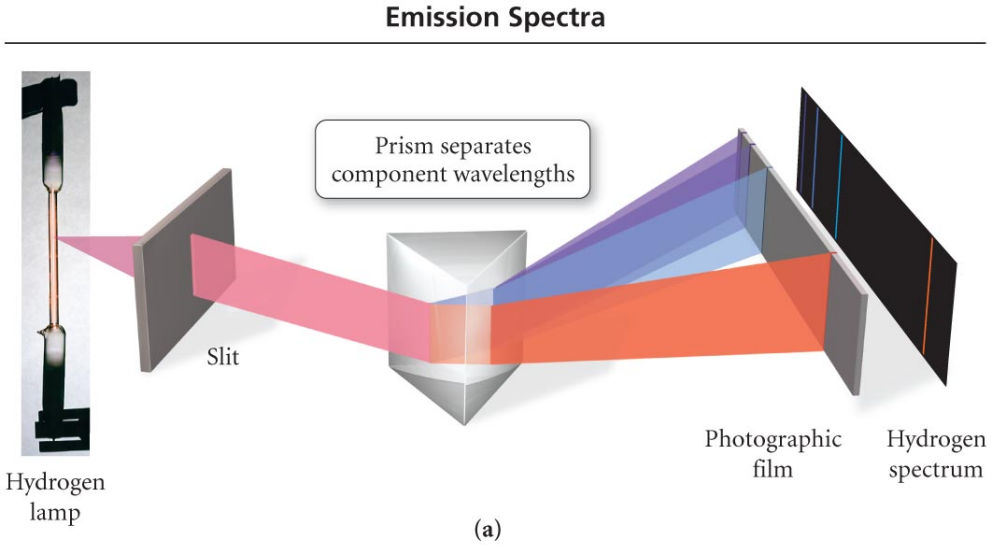
- For the photoelectric effect to occur, light must have a minimum *frequency*, called the **threshold frequency** (not a minimum *amplitude* as would have been predicted)**.**
- **Einstein** explained the photoelectric effect by proposing that light travels in packets of energy known as **photons**.
  - ➢ Light has particle as well as wave nature.

7.3 Atomic Spectra

- When gaseous elements are *excited*, (*e.g.* by heat or electricity) they emit certain frequencies of light.
- The full spectrum (rainbow) is **not** observed, but rather a series of lines at specific, discreet frequencies. Each element emits a characteristic set of frequencies.

Figure 7.11





- For **hydrogen atoms only**, the wavelengths of the lines emitted are described by the Rydberg Equation:

- n_1 and n_2 are **Quantum Numbers** ($n_2 > n_1$)
- The quantum numbers are integers (1, 2, 3…). n = 1 is the **ground** state; higher values of n are called **excited** states.

Examples:

1. Calculate the wavelength (in nm) of the photon emitted by a hydrogen atom when an electron drops from the $n = 5$ state to the $n = 3$ state.

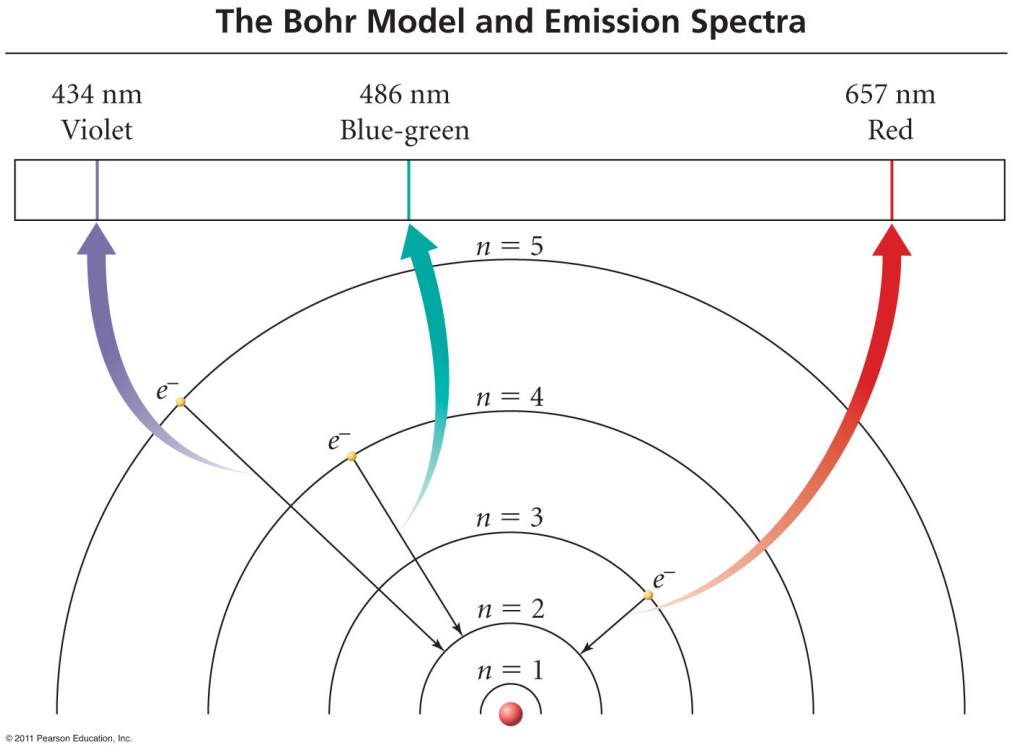
Examples:

2. What is the energy difference between the $n = 5$ and $n = 3$ states in the hydrogen atom?

Bohr Model of the Atom

- Niels Bohr proposed that the electron in the hydrogen atom "orbits" a central nucleus.
- The electron only has certain allowed orbits, ***n*** where n = 1, 2, 3… Increasing n corresponds to increasing distance from the nucleus, and higher energy.
- An electron can "jump" from one orbit to another.

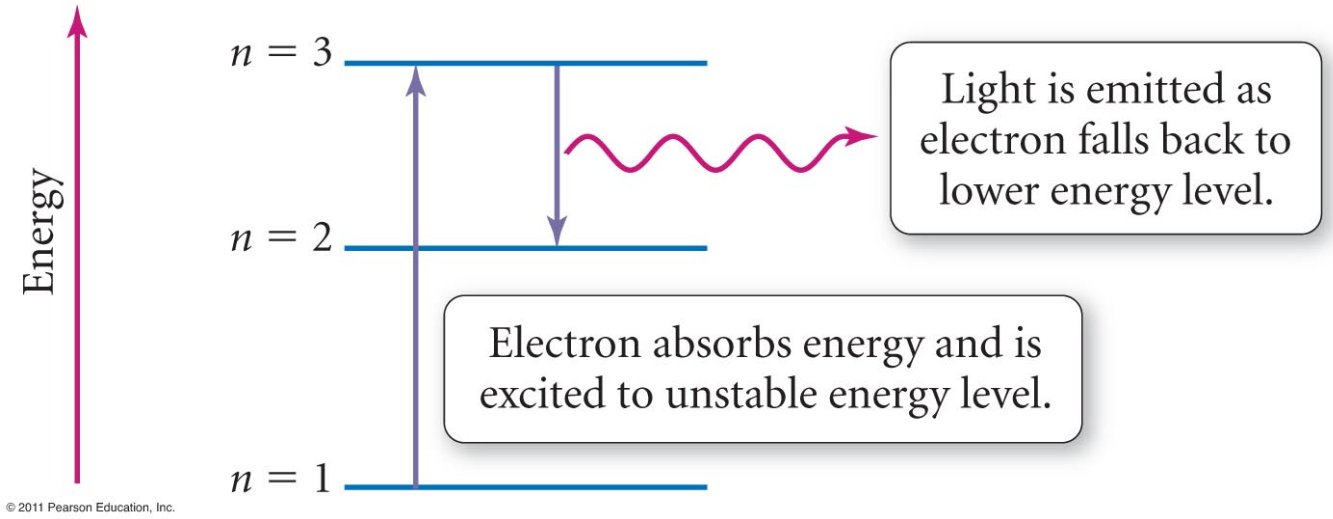
Figure 7.12



Bohr Model of the Atom (contd.)

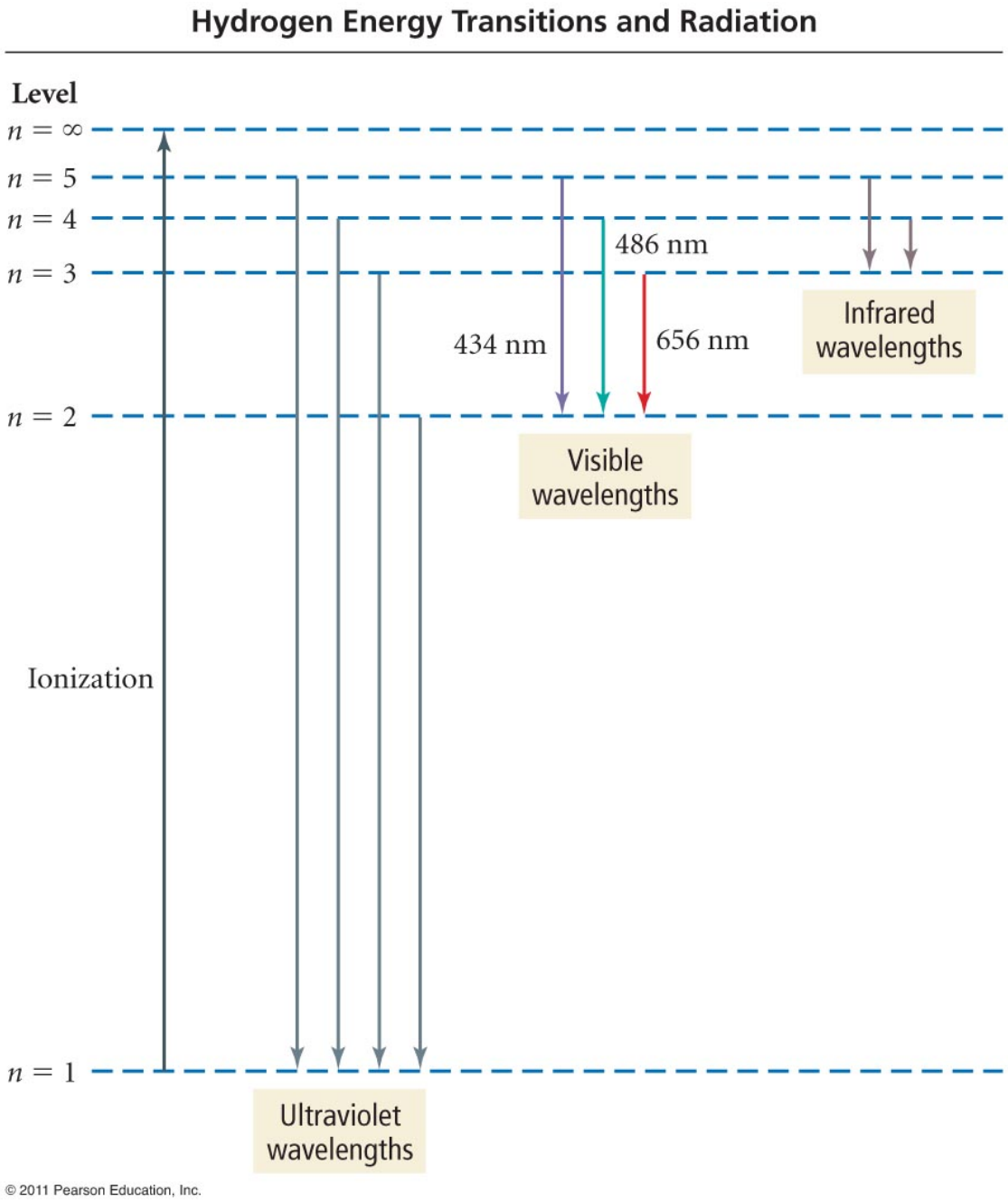
- On transferring between energy levels:
  - Light is **absorbed** when n increases.
  - Light is **emitted** when n decreases.

Figure 7.21 Excitation and Radiation



- So since only certain energy levels are possible, only certain wavelengths of light can be emitted. This is why atomic spectra consist of discreet lines.

Figure 7.22


Hydrogen Energy Transitions and Radiation
Level
n = ∞
n = 5
n = 4
n = 3
n = 2
n = 1
Ionization
434 nm
486 nm
656 nm
Infrared wavelengths
Visible wavelengths
Ultraviolet wavelengths
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Bohr Model of the Atom (contd.)

- The Bohr Model led to the concept that **energy is quantized**: An atom cannot have a continuous range of energies (like stairs rather than a ramp).
- The Bohr model of the atom only works for systems with just one electron (*e.g.* hydrogen atoms), but the concept that atoms have discreet, quantized energy levels is valid for all elements.