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Chapter 9

Atomic Physics

Atomic Physics

- Classical physics (Newtonian physics) – development of physics prior to around 1900
- Classical physics was generally concerned with macrocosm – the description & explanation of large-scale phenomena
  - cannon balls, planets, wave motion, sound, optics, electromagnetism
- Modern physics (after about 1900) – concerned with the microscopic world – microcosm – the subatomic world is difficult to describe with classical physics
- This chapter deals with only a part of modern physics called Atomic Physics – dealing with electrons in the atom.

Audio Link

Early Concepts of the Atom

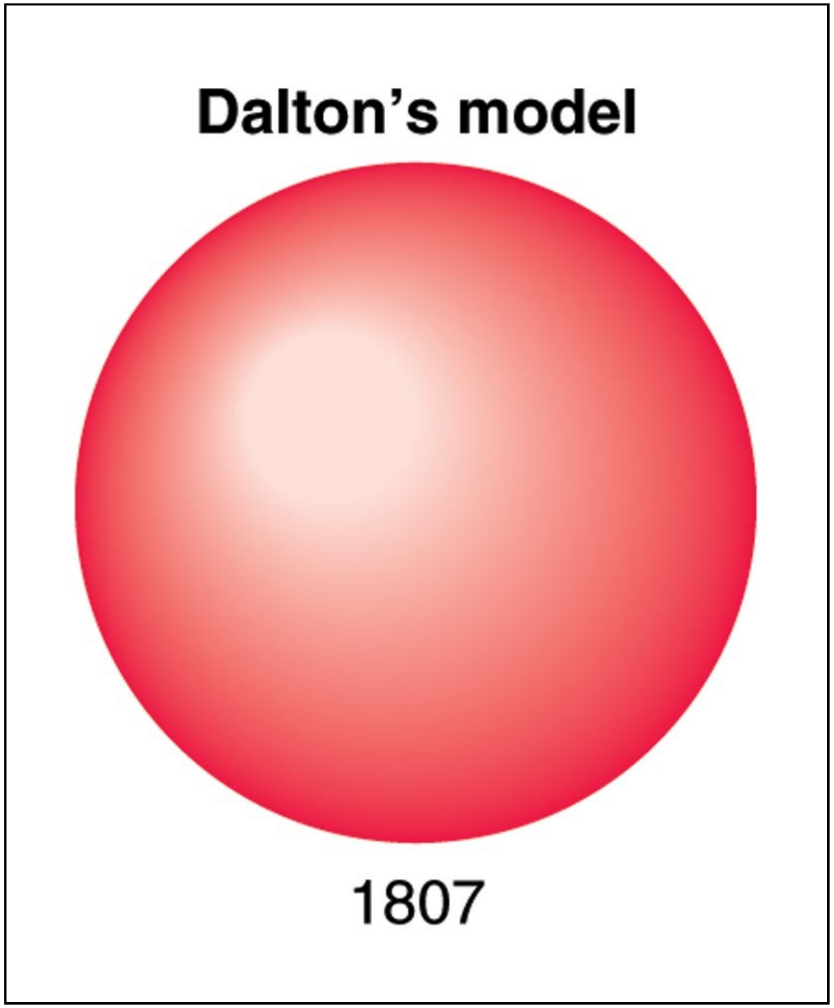
- Greek Philosophers (400 B.C.) debated whether matter was *continuous* or *discrete*, but could prove neither.
  - Continuous – could be divided indefinitely
  - Discrete – ultimate indivisible particle
  - Most (including Aristotle) agreed with the continuous theory.
- The *continuous* model of matter prevailed for 2200 years, until 1807.

Dalton’s Model – “The Billiard Ball Model”

- In 1807 John Dalton presented evidence that matter was discrete and must exist as particles.
- Dalton’s major hypothesis stated that:
- Each chemical element is composed of small indivisible particles called atoms,
  - identical for each element but different from atoms of other elements
- Essentially these particles are featureless spheres of uniform density.

Dalton’s Model

- Dalton’s 1807 “billiard ball model” pictured the atom as a tiny indivisible, uniformly dense, solid sphere.



Thomson – “Plum Pudding Model”

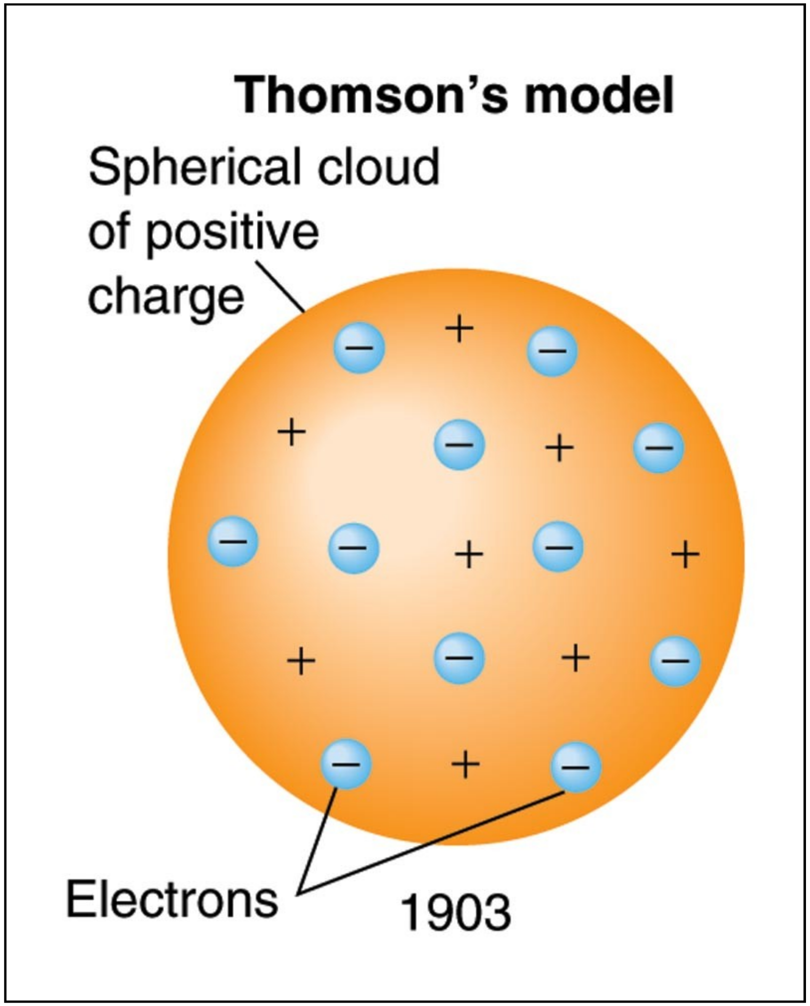
- In 1903 J.J. Thomson discovered the electron.
- Further experiments by Thomson and others showed that an electron has a mass of 9.11 x 10^{-31} kg and a charge of –1.60 x 10^{-19} C.
- Thomson produced ’rays’ using several different gas types in cathode-ray tubes.
  - He noted that these rays were deflected by electric and magnetic fields.
- Thomson concluded that this ray consisted of negative particles (now called electrons.)

Thomson – “Plum Pudding Model” (cont.)

- Identical electrons were produced no matter what gas was in the tube.
- Therefore he concluded that atoms of all types contained ’electrons.’
- Since atoms as a whole are electrically neutral, some other part of the atom must be positive.
- Thomson concluded that the electrons were stuck randomly in an otherwise homogeneous mass of positively charged “pudding.”

Thompson's Model

- Thomson's 1903 "plum pudding model" conceived the atom as a sphere of positive charge in which negatively charged electrons were embedded.

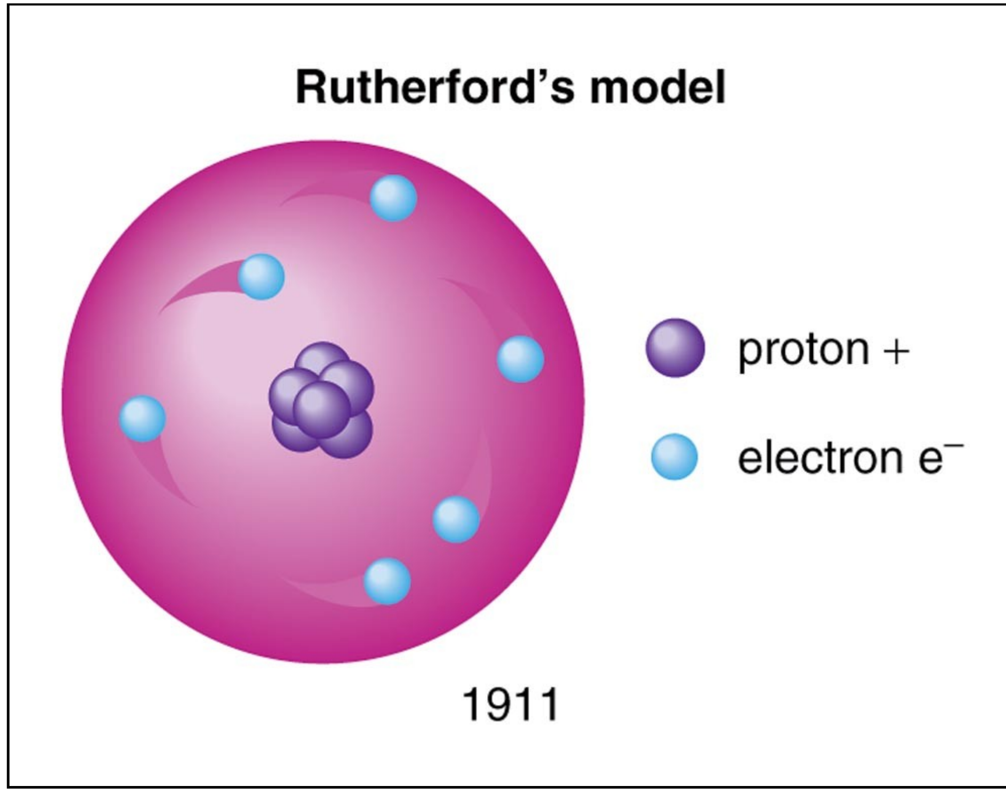


Ernest Rutherford's Model

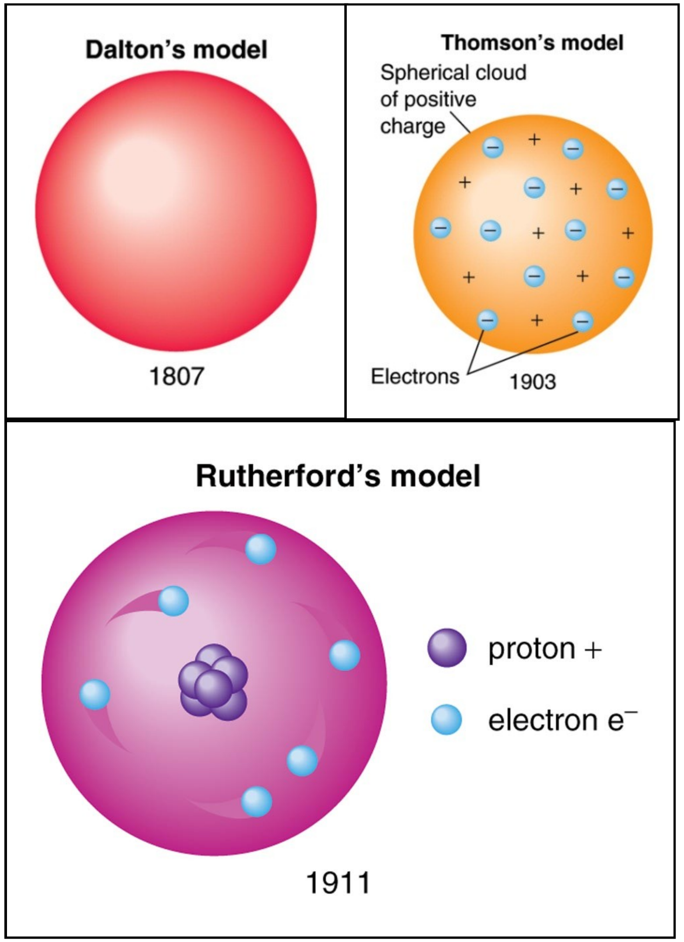
- In 1911 Rutherford discovered that 99.97% of the mass of an atom was concentrated in a tiny core, or nucleus.
- Rutherford's model envisioned the electrons as circulating in some way around a positively charged core.

Rutherford’s Model

- Rutherford’s 1911 “nuclear model” envisioned the atom as having a dense center of positive charge (the nucleus) around which the electrons orbited.



Evolution of the Atomic Models 1807 - 1911

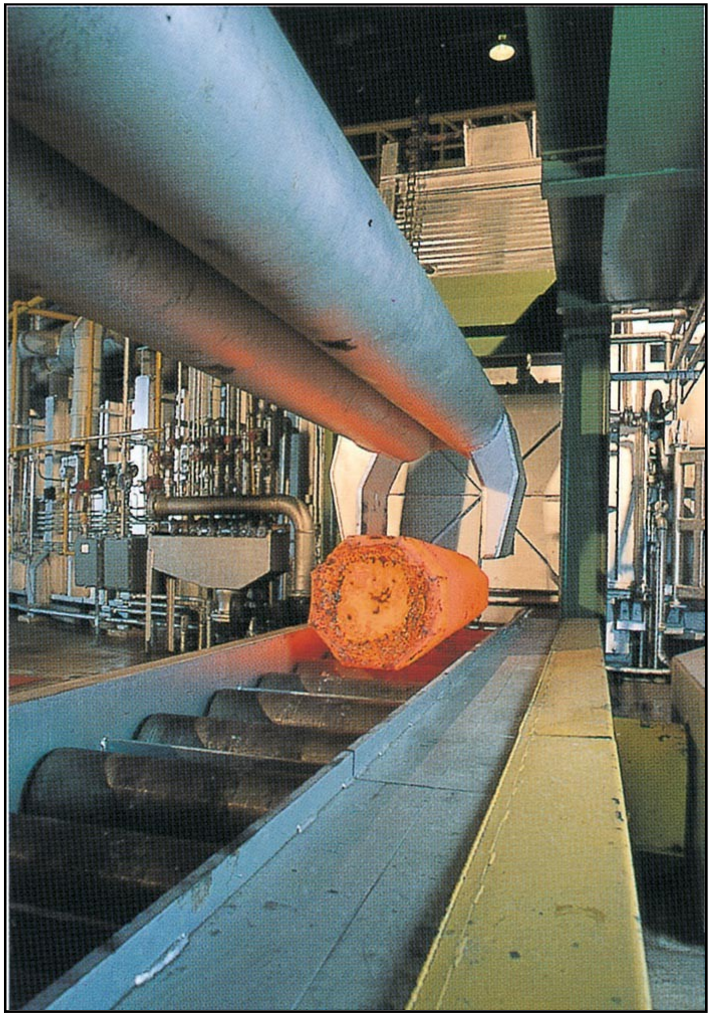


Classical Wave Theory of Light

- Scientists have known for many centuries that very hot (or incandescent) solids emit visible light
  - Iron may become “red” hot or even “white” hot, with increasing temperature
  - The light that common light bulbs give off is due to the incandescence of the tungsten filament
- This increase in emitted light frequency is expected because as the temperature increases the greater the electron vibrations and $\therefore$ the higher the frequency of the emitted radiation

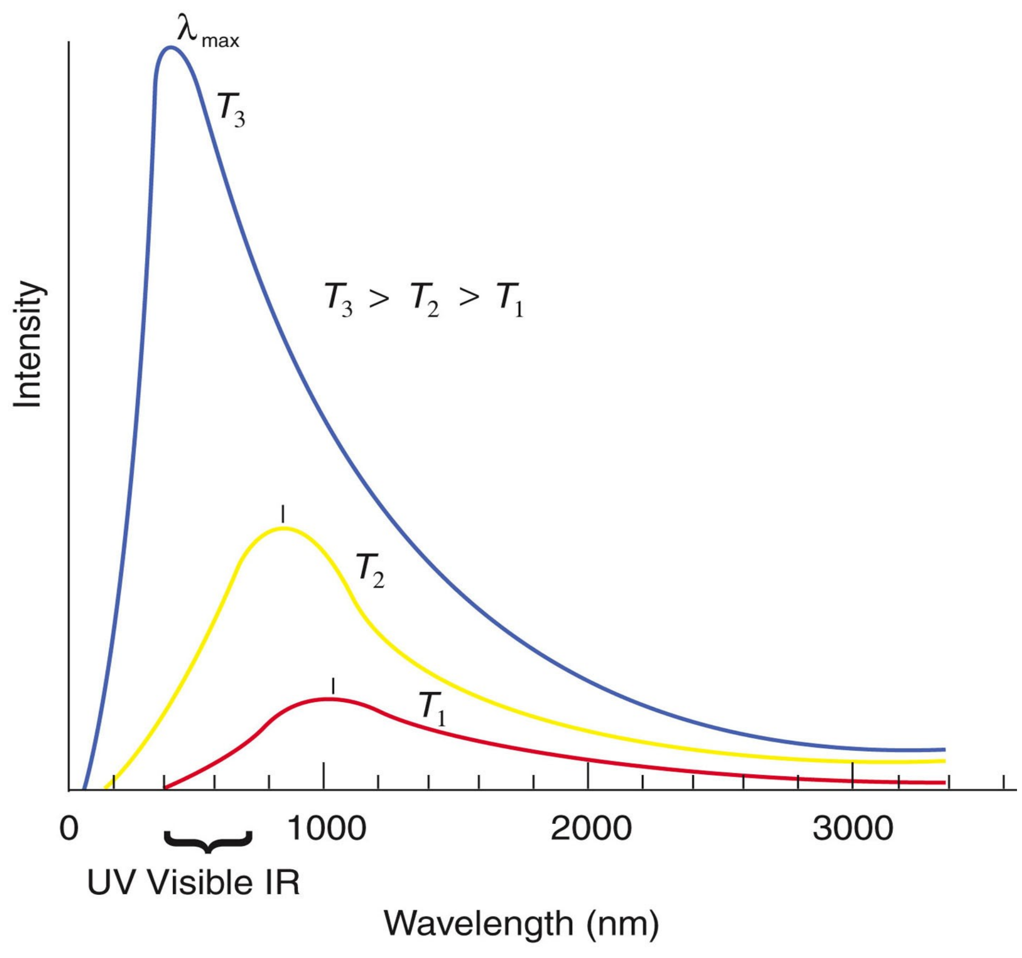
Red-Hot Steel

- The radiation component of maximum intensity determines a hot solid's color.



Thermal Radiation

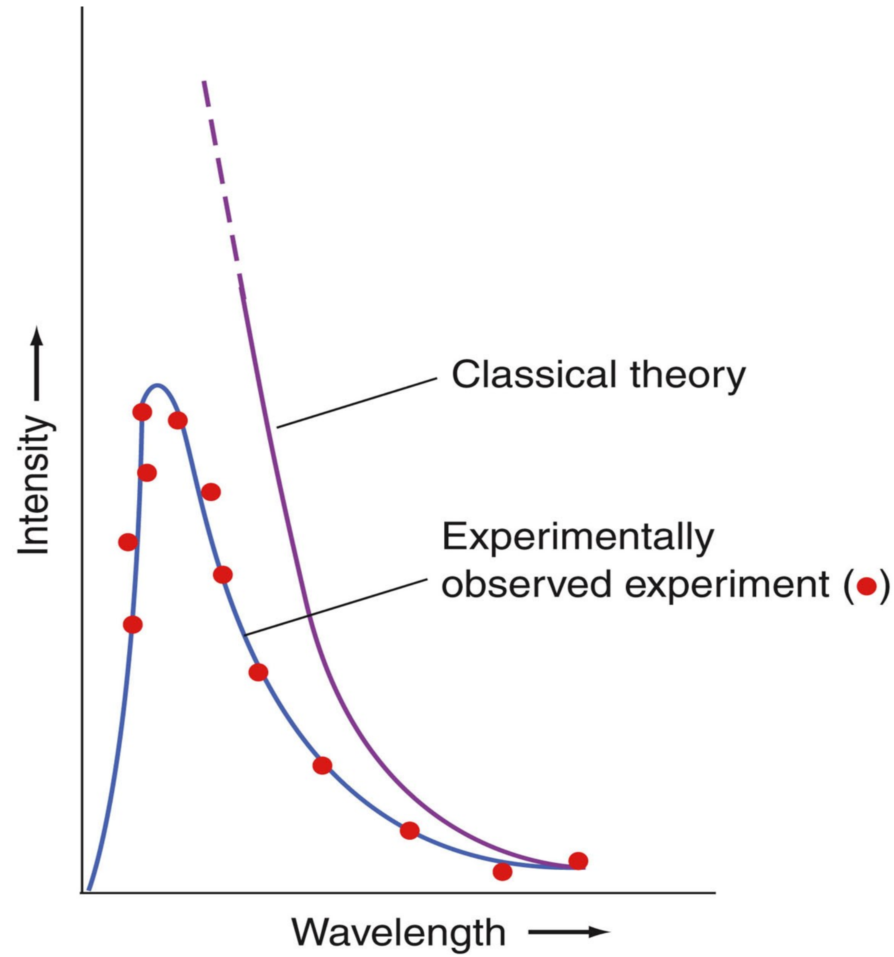
- As the temperature increases the peak of maximum intensity shifts to higher frequency – it goes from red to orange to white hot. Wave theory correctly predicts this.



(a)

Classical Wave Theory

- According to classical wave theory, $I \propto f^2$. This means that I should increase rapidly (exponentially) as f increases.
- This is NOT what is actually observed.



(b)

The Ultraviolet Catastrophe

- Since classical wave could not explain why the relationship $I \alpha f^2$ is not true, this dilemma was coined the “ultraviolet catastrophe”
  - “Ultraviolet” because the relationship broke down at high frequencies.
  - And “catastrophe” because the predicted energy intensity fell well short of expectations.
- The problem was resolved in 1900 by Max Planck, a German physicist

Max Planck (1858-1947)

- In 1900 Planck introduced the idea of a quantum – an oscillating electron can only have discrete, or specific amounts of energy
- Planck also said that this amount of energy (E) depends on its frequency (f)
- Energy = Planck’s constant x frequency ($E = hf$)
- This concept by Planck took the first step toward quantum physics

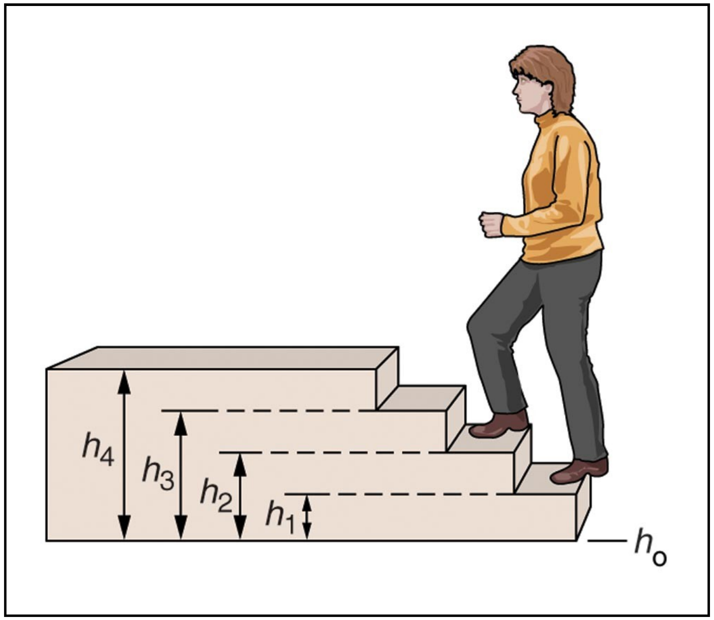
Quantum Theory

- Planck's hypothesis correctly accounted for the observed intensity with respect to the frequency squared
- Therefore Planck introduce the idea of the "quantum" – a discrete amount of energy
  - Like a "packet" of energy
- Similar to the potential energy of a person on a staircase – they can only have specific potential-energy values, determined by the height of each stair

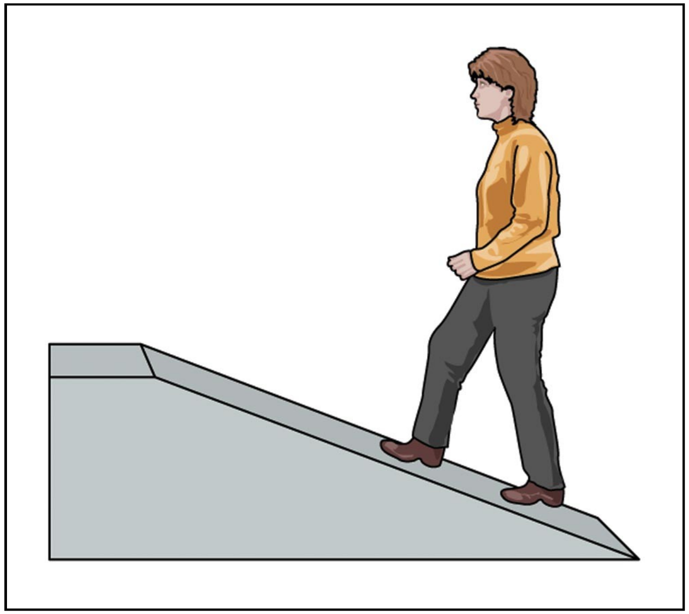
Concept of Quantized Energy

four specific potential energy values

Quantized Energy



Continuous Energy



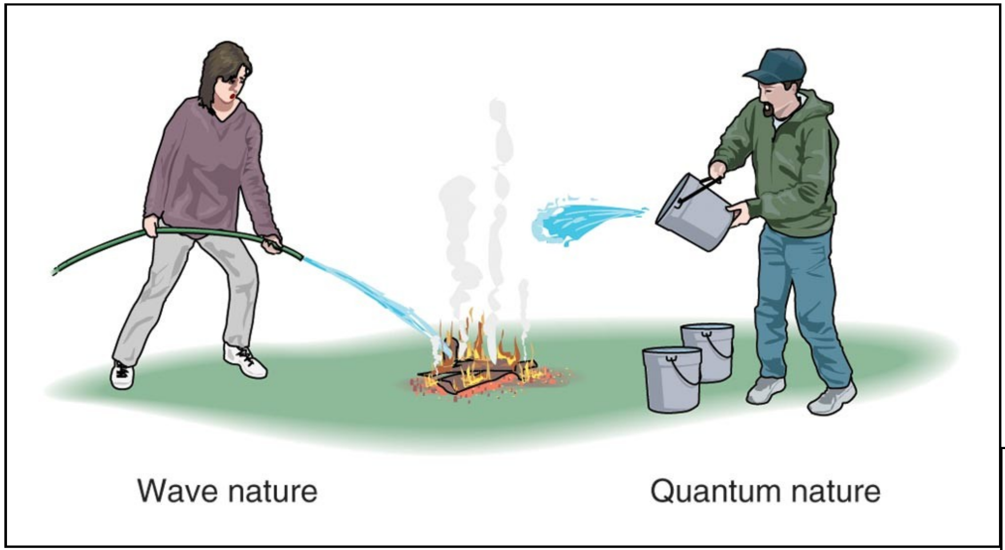
Photoelectric Effect

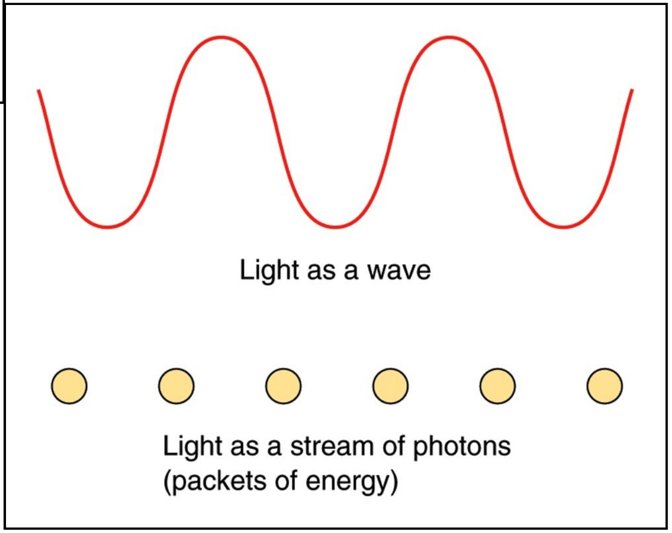
- Scientists noticed that certain metals emitted electrons when exposed to light– The photoelectric effect
- This direct conversion of light into electrical energy is the basis for photocells
  - Automatic door openers, light meters, solar energy applications
- Once again classical wave theory could not explain the photoelectric effect

Photoelectric Effect Solved by Einstein using Planck's Hypothesis

- In classical wave theory it should take an appreciable amount of time to cause an electron to be emitted
- But … electrons flow is almost immediate when exposed to light
- Thereby indicating that light consists of "particles" or "packets" of energy
- Einstein called these packets of energy "photons"

Wave Model – continuous flow of energy
Quantum Model – packets of energy





Photoelectric Effect

- In addition, it was shown that the higher the frequency the greater the energy
  - For example, blue light has a higher frequency than red light and therefore have more energy than photons of red light.
- In the following two examples you will see that Planck’s Equation correctly predict the relative energy levels of red and blue light

Determining Photon Energy
Example of how photon energy is determined

- *Find the energy in joules of the photons of red light of frequency 5.00 x 10^{14} Hz (cycles/second)*
- You are given: f and h (Planck's constant)
- Use Planck's equation $E = hf$
- $E = hf$ = (6.63 x 10^{-34}J s)(5.00 x 10^{14}/sec)
- = 33.15 x 10^{-20} J

Audio Link

Determining Photon Energy

Another example of how photon energy is determined

- *Find the energy in joules of the photons of blue light of frequency 7.50 x 10^{14} Hz (cycles/second)*
- You are given: f and *h* (Planck's constant)
- Use Planck's Equation $E = hf$
- $E = hf$ = (6.63 x 10^{-34}J s)(7.50 x 10^{14}/sec)
- = 49.73 x 10^{-20} J
- ** Note the blue light has more energy than red light

The Dual Nature of Light

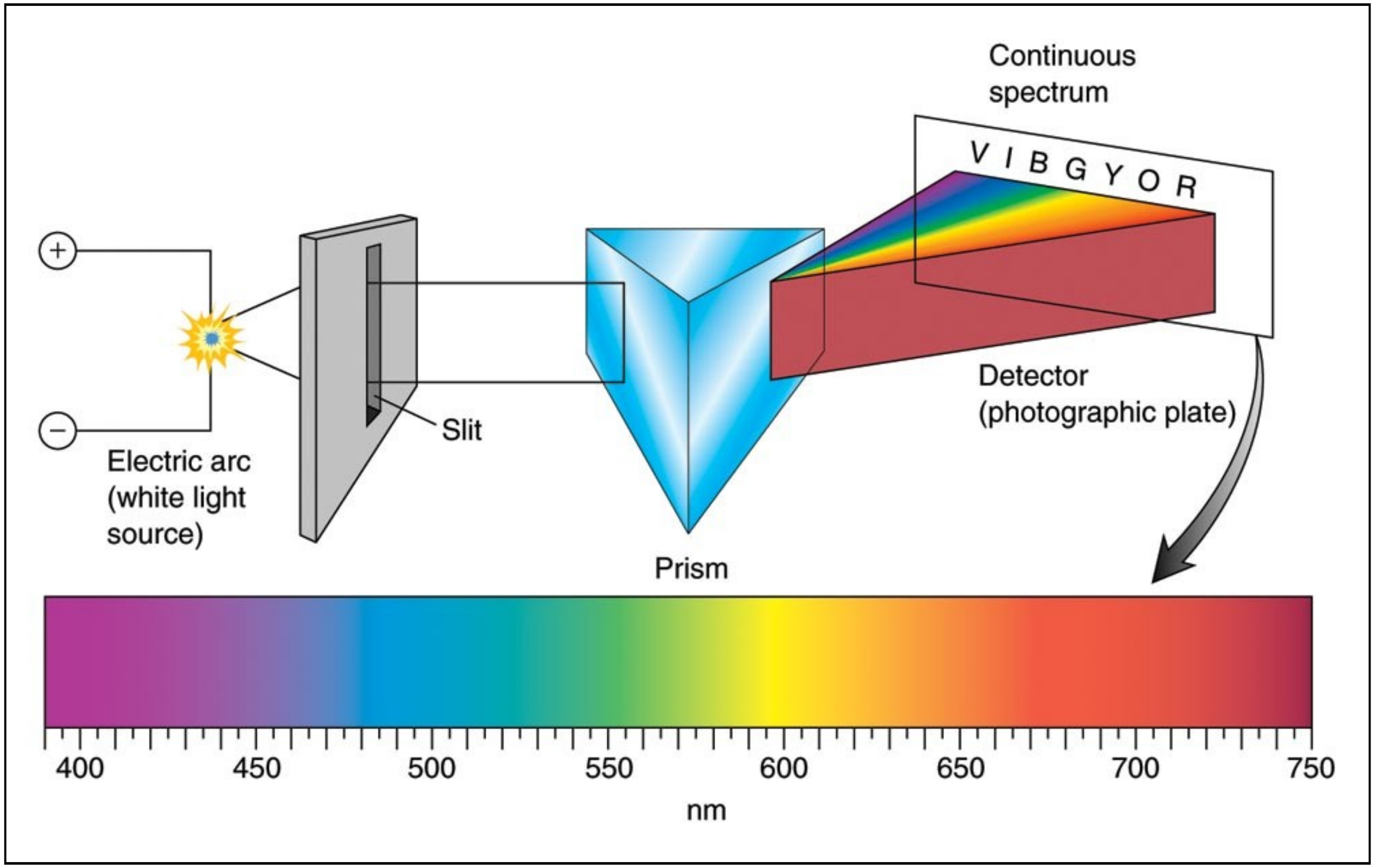
- To explain various phenomena, light sometimes must be described as a wave and sometimes as a particle.
- Therefore, in a specific experiment, scientists use <u>whichever</u> model (wave or particle theory) of light works!!
- Apparently light is not exactly a wave or a particle, but has characteristics of both
- In the microscopic world our macroscopic analogies may not adequately fit

Two Types of Spectra

- Recall from chapter 7 that white light can be dispersed into a spectrum of colors by a prism
  - Due to differences in refraction for the specific wavelengths
- In the late 1800's experimental work with gas-discharge tubes revealed two other types of spectra
  - Line emission spectra displayed only bright spectral lines of certain frequencies
  - Line absorption spectra displays dark lines of missing colors

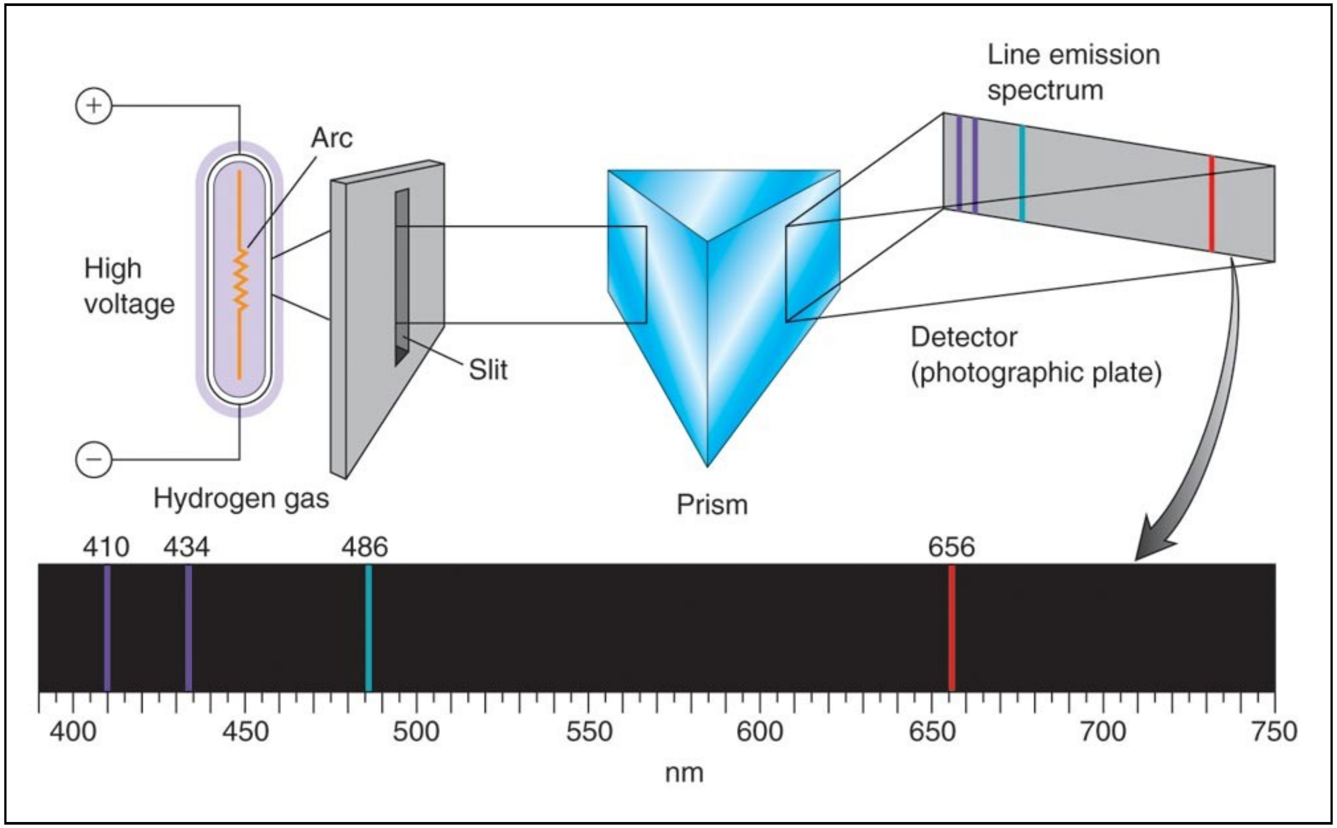
Continuous Spectrum of Visible Light

Light of all colors is observed



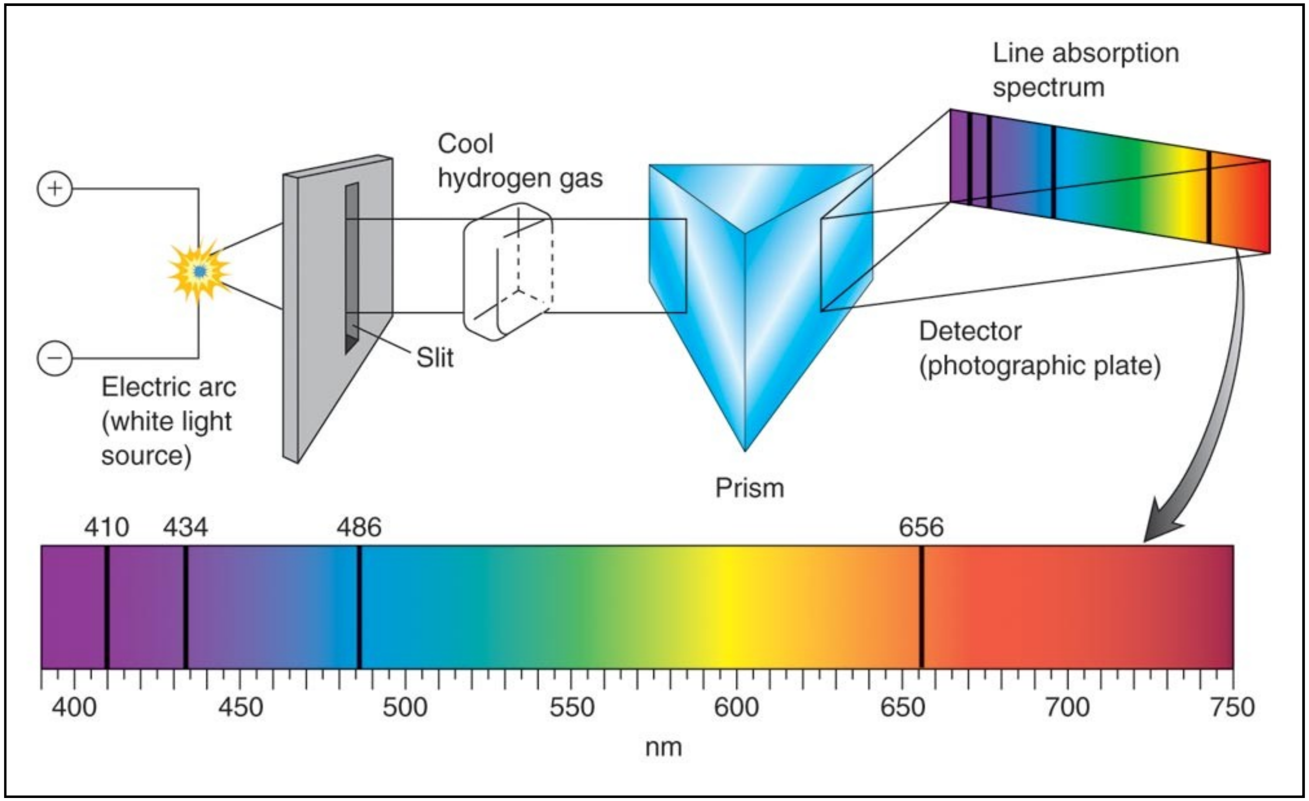
Line Emission Spectrum for Hydrogen

- When light from a gas-discharge tube is analyzed only spectral lines of certain frequencies are found



Line Absorption Spectrum for Hydrogen

- Results in dark lines (same as the bright lines of the line emission spectrum) of missing colors.



Spectra & the Bohr Model

- Spectroscopists did not initially understand why only discrete, and characteristic wavelengths of light were
  - Emitted in a line emission spectrum, and
  - Omitted in a line absorption spectrum
- In 1913 an explanation of the observed spectral line phenomena was advanced by the Danish physicist Niels Bohr

Bohr and the Hydrogen Atom

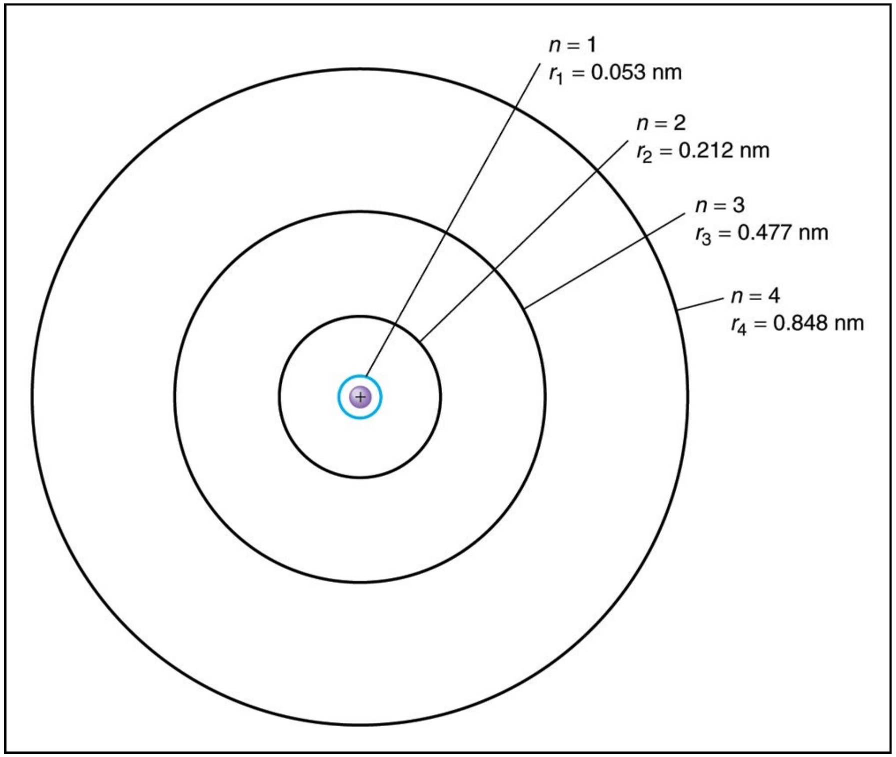
- Bohr decided to study the hydrogen atom because it is the simplest atom
  - One single electron “orbiting” a single proton
- As most scientists before him, Bohr assumed that the electron revolved around the nuclear proton – but…
- Bohr correctly reasoned that the characteristic (and repeatable) line spectra were the result of a “quantum effect”

Bohr and the Hydrogen Atom

- Bohr predicted that the single hydrogen electron would only be found in discrete orbits with particular radii
  - Bohr's possible electron orbits were given whole-number designations, n = 1, 2, 3, …
  - "n" is called the principal quantum number
  - The lowest n-value, (n = 1) has the smallest radius

Bohr Electron Orbits

- Each possible electron orbit is characterized by a quantum number.
- Distances from the nucleus are given in nanometers.

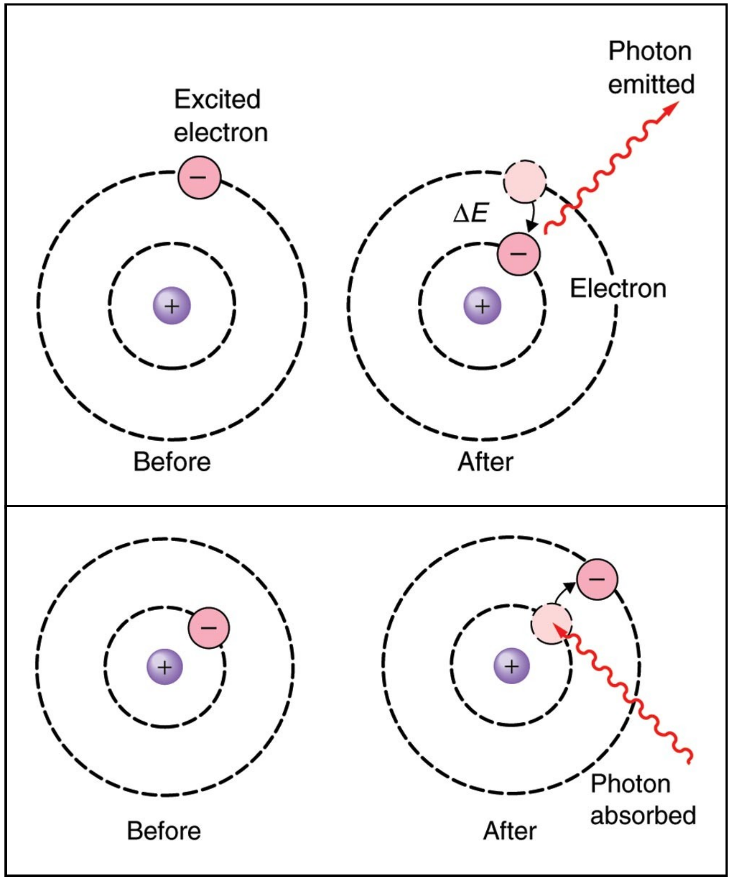


Bohr and the Hydrogen Atom

- Classical atomic theory indicated that an accelerating electron should continuously radiate energy
  - But this is not the case
  - If an electron continuously lost energy, it would soon spiral into the nucleus
- Bohr once again correctly hypothesized that the hydrogen electron only radiates/absorbs energy when it makes an quantum jump or transition to another orbit

Photon Emission and Absorption

- A transition to a lower energy level results in the emission of a photon.
- A transition to a higher energy level results in the absorption of a photon.



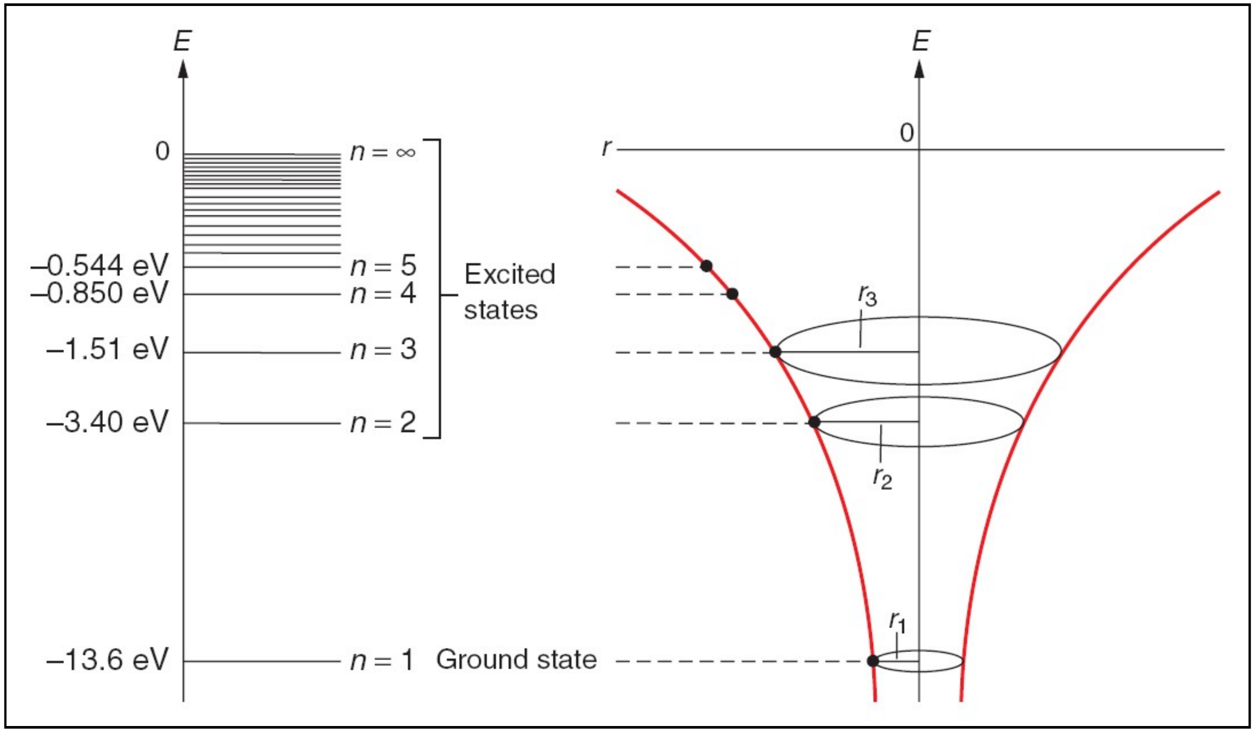
The Bohr Model

- According to the Bohr model the “allowed orbits” of the hydrogen electron are called energy states or energy levels
  - Each of these energy levels correspond to a specific orbit and principal quantum number
- In the hydrogen atom, the electron is normally at $n = 1$ or the ground state
- The energy levels above the ground state ($n = 2, 3, 4, \ldots$) are called excited states

Orbits and Energy Levels of the Hydrogen Atom

Bohr theory predicts that the hydrogen electron can only occupy discrete radii.

- Note that the energy levels are <u>not</u> evenly spaced.



The Bohr Model

- If enough energy is applied, the electron will no longer be bound to the nucleus and the atom is *ionized*
- As a result of the mathematical development of Bohr’s theory, scientists are able to predict the radii and energies of the allowed orbits
- For hydrogen, the radius of a particular orbit can be expressed as
  - $r_n = 0.053\, n^2$ nm
    - n = principal quantum number of an orbit
    - r = orbit radius

Confidence Exercise

Determining the Radius of an Orbit in a Hydrogen Atom

- Determine the radius in nm of the second orbit (n = 2, the first excited state) in a hydrogen atom
- Solution:
- Use equation 9.2 → $r_n = 0.053\ n^2$ nm
- $n = 2$
- $r_1 = 0.053\ (2)^2$ nm = <u>0.212 nm</u>
- *Same value as Table 9.1!*

Energy of a Hydrogen Electron

- The total energy of the hydrogen electron in an allowed orbit is given by the following equation:
- $E_n = -13.60/n^2$ eV (eV = electron volts)
  - The negative sign means it is in a potential well
- The ground state energy value for the hydrogen electron is –13.60 eV
  - $\therefore$ it takes 13.60 eV to ionize a hydrogen atom
  - the hydrogen electron's <u>binding energy</u> is 13.60 eV
- Note that as the n increases the energy levels become closer together

Problem Example
Determining the Energy of an Orbit in the Hydrogen Atom

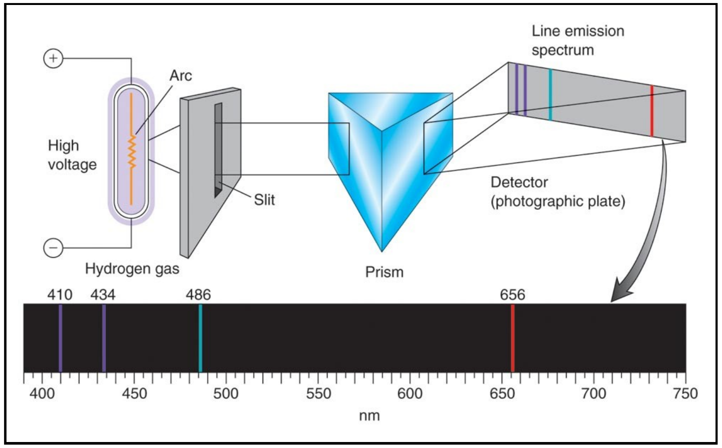
- Determine the energy of an electron in the first orbit (n = 1, the ground state) in a hydrogen atom
- Solution:
- Use equation 9.3 → $E_n = -13.60/n^2\ \mathrm{eV}$
- n = 1
- $E_n = -13.60/(1)^2\ \mathrm{eV} = \underline{-13.60\ \mathrm{eV}}$
- *Same value as Table 9.1!*

Confidence Exercise
Determining the Energy of an Orbit in the Hydrogen Atom

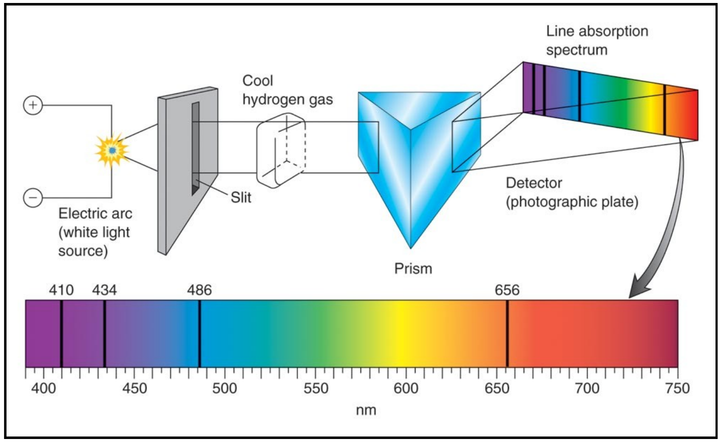
- Determine the energy of an electron in the first orbit (n = 2, the first excited state) in a hydrogen atom
- Solution:
- Use equation 9.3 → $E_n = -13.60/n^2$ eV
- n = 2
- $E_n = -13.60/(2)^2$ eV = $\underline{-3.40 \text{ eV}}$
- *Same value as Table 9.1!*

Explanation of Discrete Line Spectra

- Recall that Bohr was trying to explain the discrete line spectra as exhibited in the -
  - Line Emission & Line Absorption spectrum
  - Note that the observed and omitted spectra coincide!



Line Emission Spectra



Line Absorption Spectra

Explanation of Discrete Line Spectra

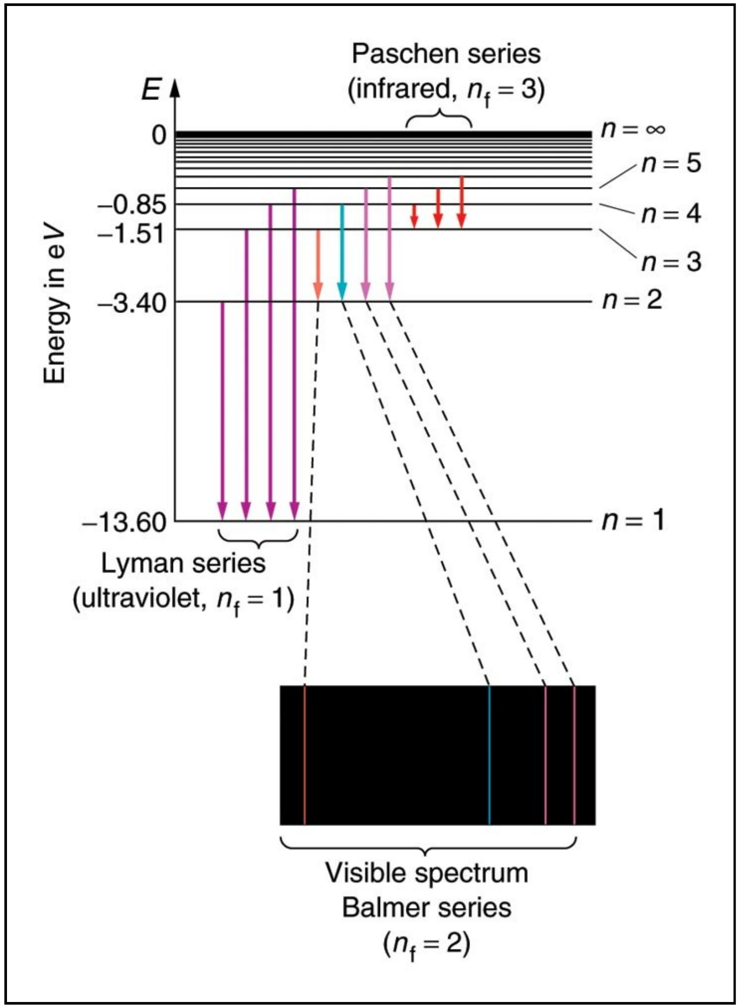
- The hydrogen line emission spectrum results from the emission of energy as the electron de-excites
  - Drops to a lower orbit and emits a photon
  - $E_{total} = E_{n_{initial}} = E_{n_f} + E_{photon}$
- The hydrogen line absorption spectrum results from the absorption of energy as the electron is excited
  - Jumps to a higher orbit and absorbs a photon

Bohr Hypothesis Correctly Predicts Line Spectra

- The dark lines in the hydrogen line absorption spectrum exactly matches up with the bright lines in the hydrogen line emission spectrum
- Therefore, the Bohr hypothesis correctly predicts that an excited hydrogen atom will emit/absorb light at the same discrete frequencies/amounts, depending upon whether the electron is being excited or de-excited

Spectral Lines for Hydrogen

- Transitions among discrete energy orbit levels give rise to discrete spectral lines within the UV, visible, and IR wavelengths



Quantum Effect

- Energy level arrangements are different for all of the various atoms
- Therefore every element has a characteristic and unique line emission and line absorption “fingerprints”
- In 1868 a dark line was found in the solar spectrum that was unknown at the time
  - It was correctly concluded that this line represented a new element – named helium
  - Later this element was indeed found on Earth

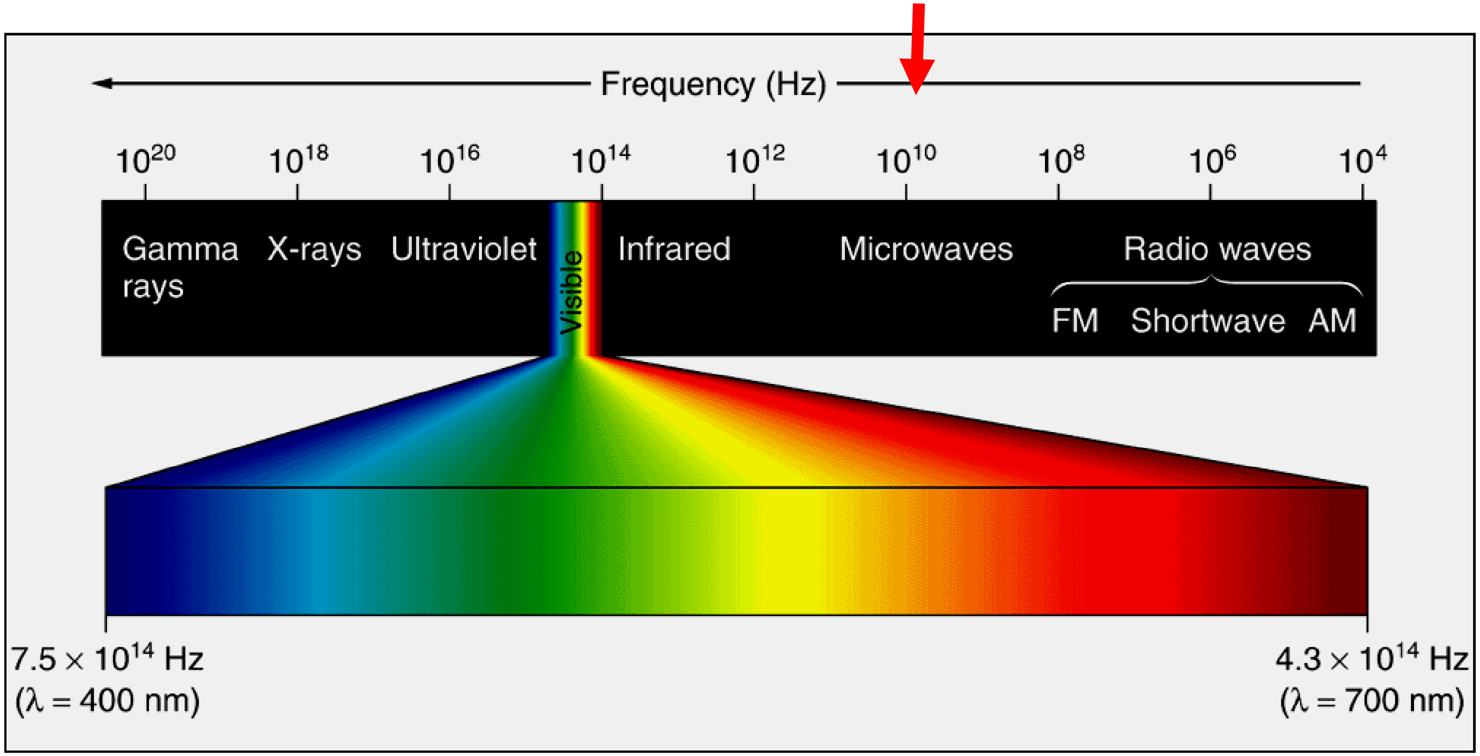
Audio Link

Molecular Spectroscopy

- Modern Physics and Chemistry actively study the energy levels of various atomic and molecular systems
  - Molecular Spectroscopy is the study of the spectra and energy levels of molecules
- As you might expect, molecules of individual substances produce unique and varied spectra
- For example, the water molecule has rotational energy levels that are affected and changed by microwaves

Microwaves

- Electromagnetic radiation that have relatively low frequencies (about $10^{10}\,\mathrm{Hz}$)



The Microwave Oven

- Because most foods contain moisture, their water molecules absorb the microwave radiation and gain energy
  - As the water molecules gain energy, they rotate more rapidly, thus heating/cooking the item
  - Fats and oils in the foods also preferentially gain energy from (are excited by) the microwaves

The Microwave Oven

- Paper/plastic/ceramic/glass dishes are not directly heated by the microwaves
  - But may be heated by contact with the food (conduction)
- The interior metal sides of the oven reflect the radiation and remain cool
- Do microwaves penetrate the food and heat it throughout?
  - Microwaves only penetrate a few centimeters and therefore they work better if the food is cut into small pieces
  - Inside of food must be heated by conduction

“Discovery” of Microwaves as a Cooking Tool

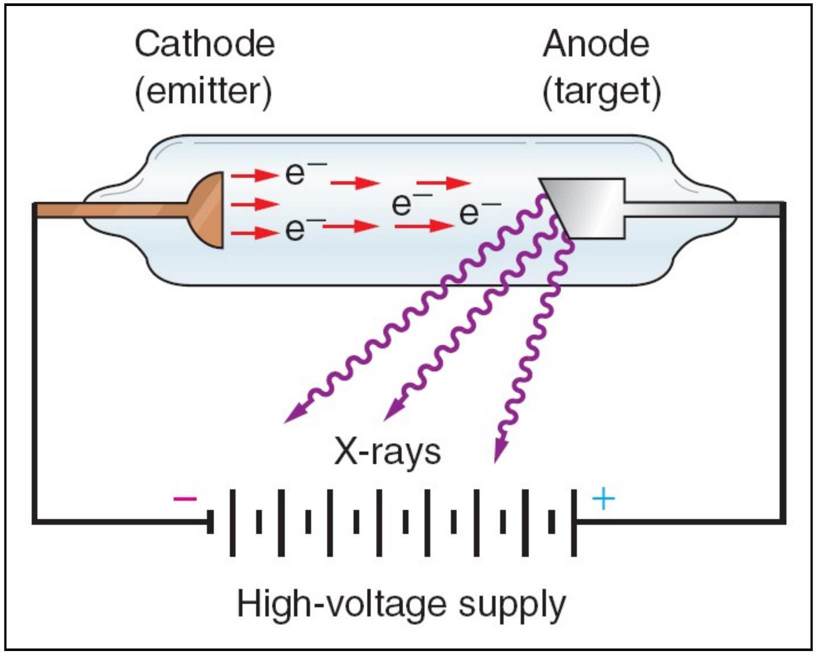
- In 1946 a Raytheon Corporation engineer, Percy Spencer, put his chocolate bar too close to a microwave source
- The chocolate bar melted of course, and …
- Within a year Raytheon introduced the first commercial microwave oven!

X-Rays

- Accidentally discovered in 1895 by the German physicist Wilhelm Roentgen
  - He noticed while working with a gas-discharge tube that a piece of fluorescent paper across the room was glowing
- Roentgen deduced that some unknown/unseen radiation from the tube was the cause
  - He called this mysterious radiation “X-radiation” because it was unknown

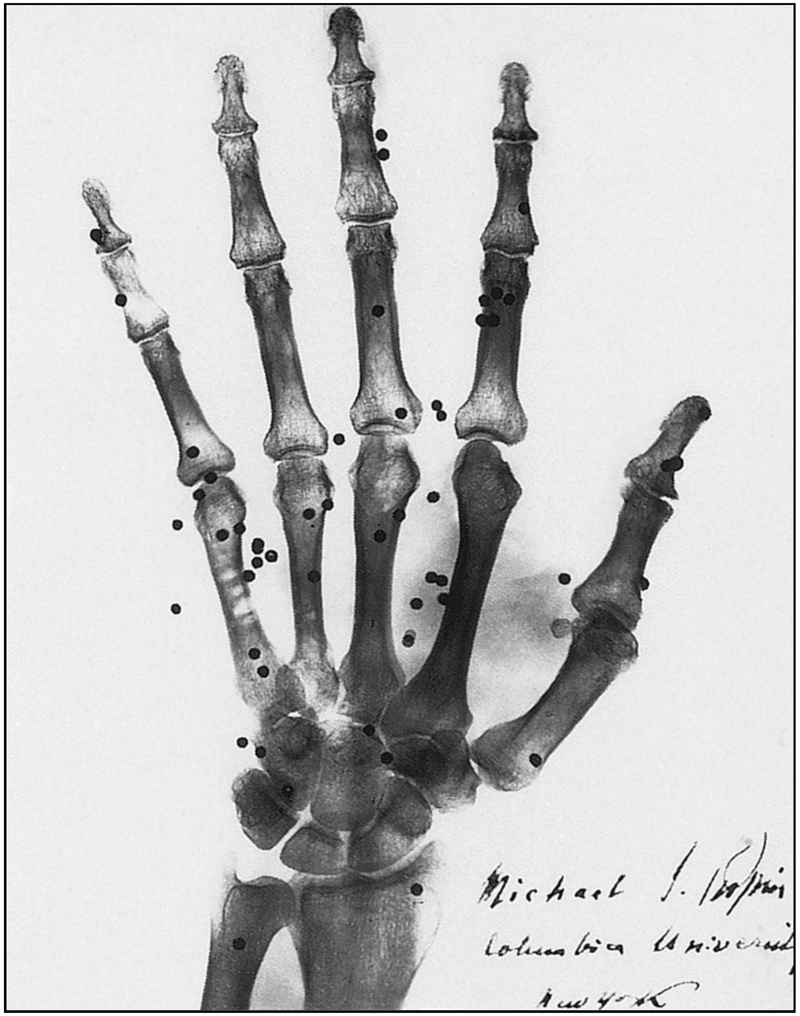
X-Ray Production

- Electrons from the cathode are accelerated toward the anode. Upon interacting with the atoms of the anode, the atoms emit energy in the form of x-rays.



Early use of X-Rays

- Within few months of their discovery, X-rays were being put to practical use.
- This is an X-ray of bird shot embedded in a hand.
- Unfortunately, much of the early use of X-rays was far too aggressive, resulting in later cancer.

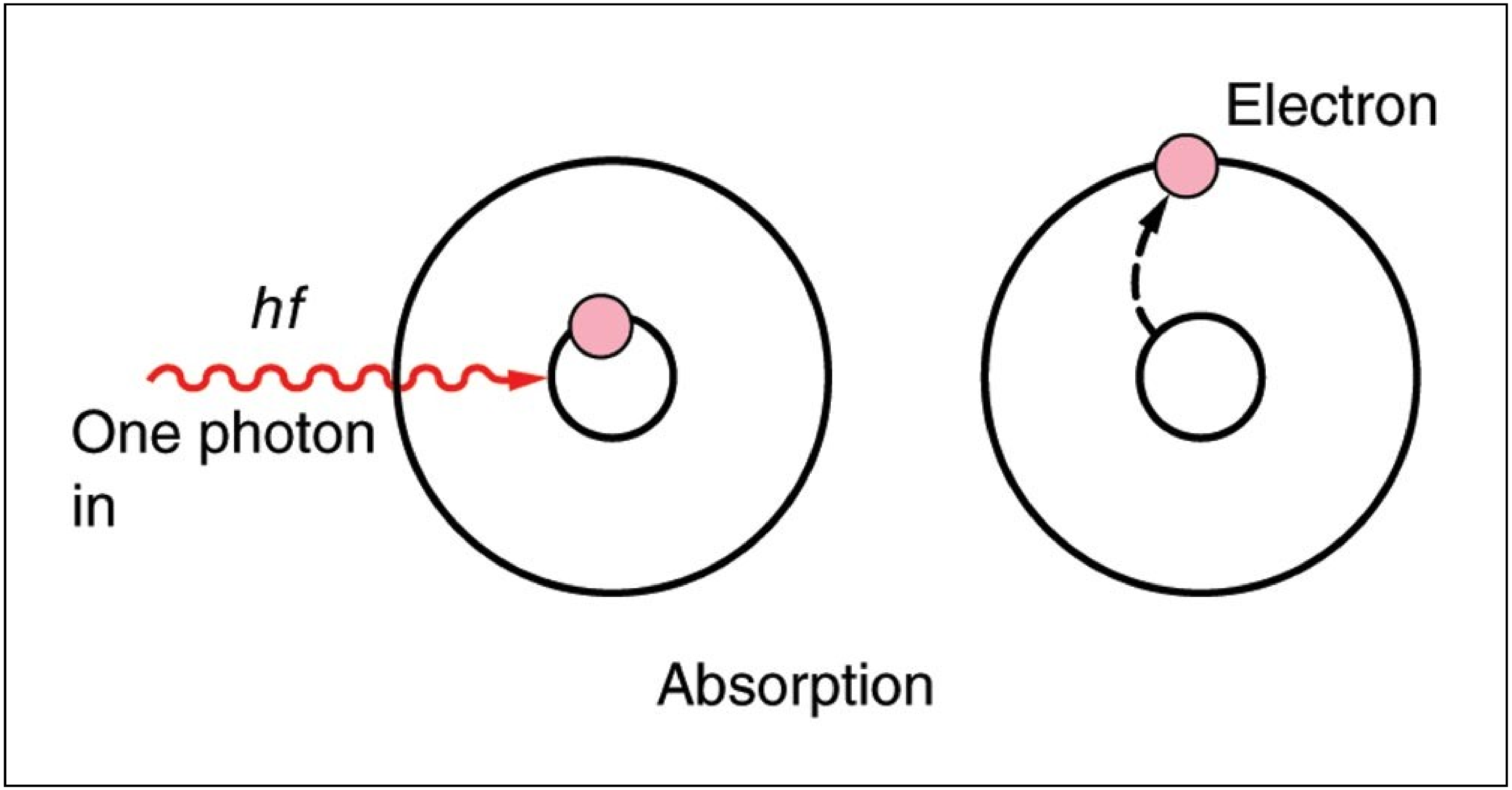


Lasers

- Unlike the accidental discovery of X-rays, the idea for the laser was initially developed from theory and only later built
- The word laser is an acronym for
  - Light Amplification by Stimulated Emission of Radiation
- Most excited atoms will immediately return to ground state, but …
- Some substances (ruby crystals, CO_2 gas, and others) have metastable excited states

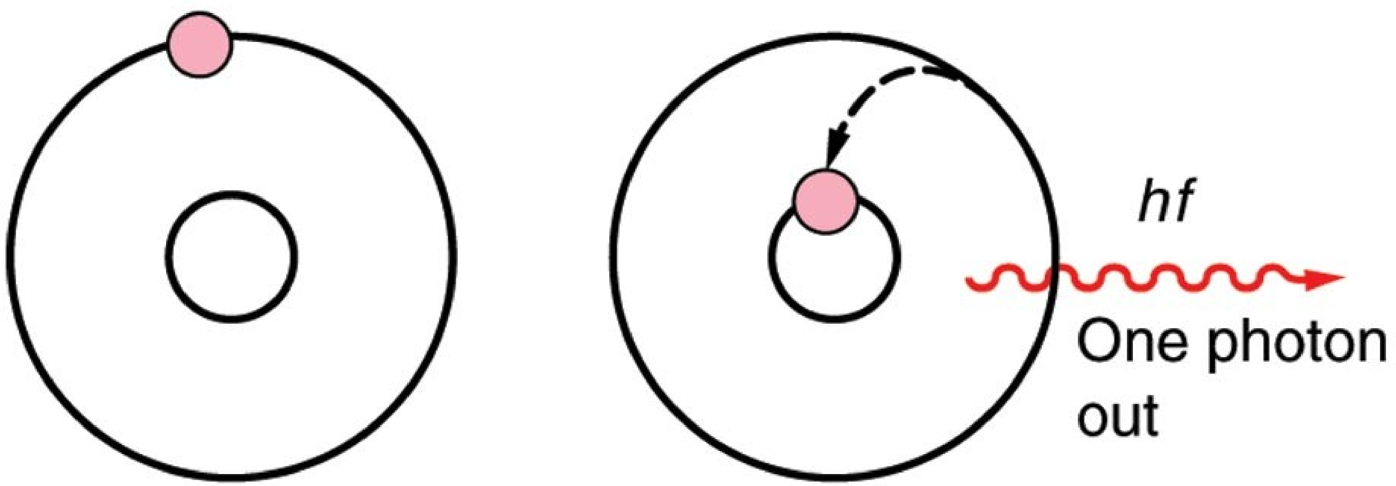
Photon Absorption

- An atom absorbs a photon and becomes excited (transition to a higher orbit)



Spontaneous Emission

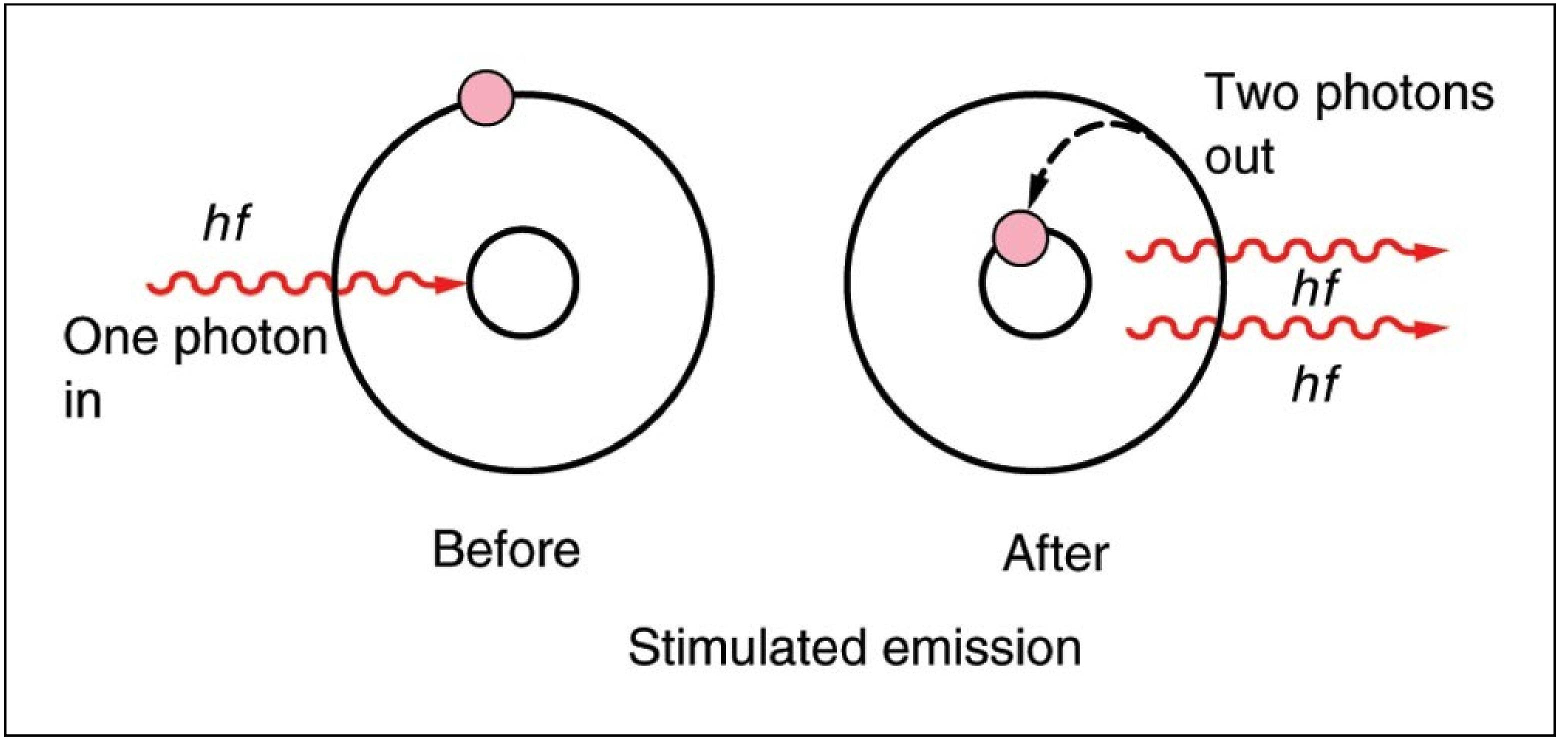
- Generally the excited atom immediately returns to ground state, emitting a photon



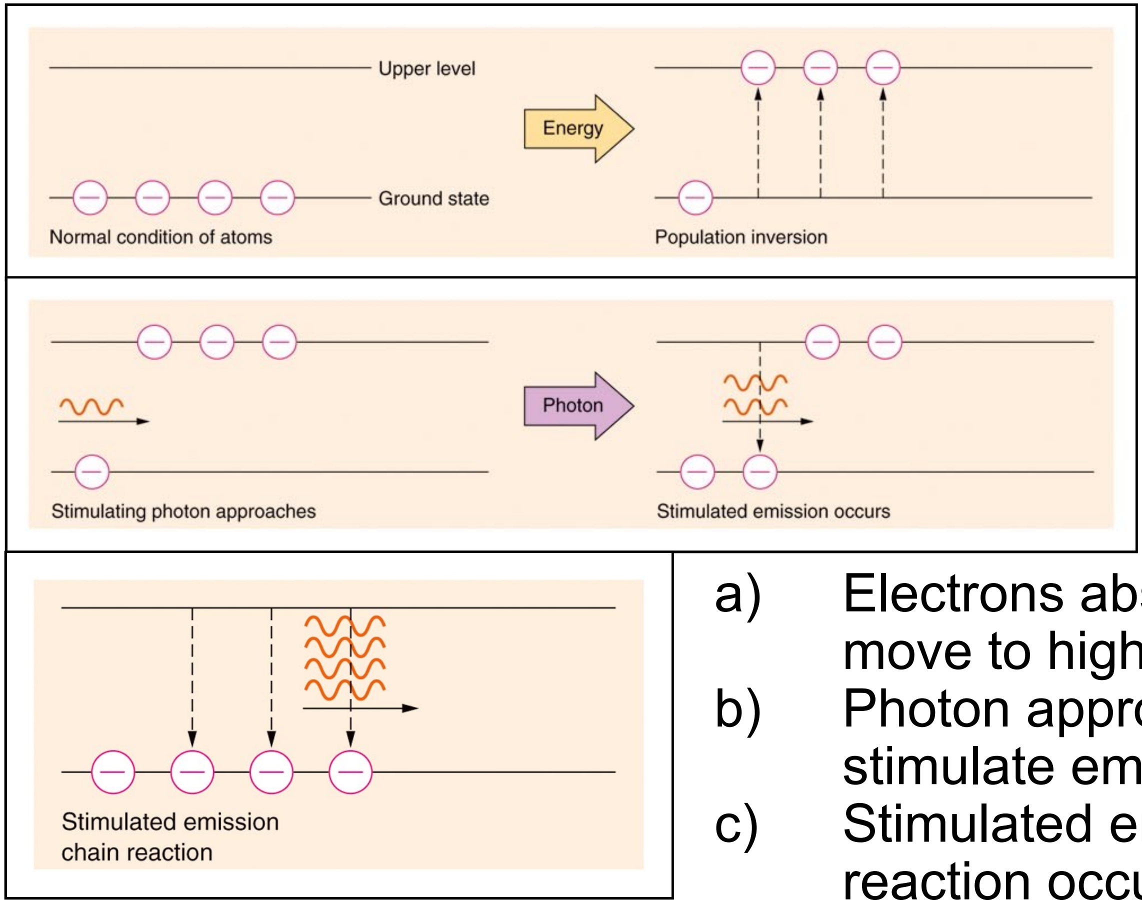
Spontaneous emission

Stimulated Emission

- Striking an excited atom with a photon of the same energy as initially absorbed will result in the emission of two photons



Stimulated Emission – the Key to the Laser



a) Electrons absorb energy and move to higher level
b) Photon approaches and stimulate emission occurs
c) Stimulated emission chain reaction occurs

Laser

- In a stimulated emission an excited atom is struck by a photon of the same energy as the allowed transition, and two photons are emitted
- The two photon are in phase and therefore constructively interfere
- The result of many stimulated emissions and reflections in a laser tube is a narrow, intense beam of laser light
  - The beam consists of the same energy and wavelength (monochromatic)

Laser Uses

- Very accurate measurements can be made by reflecting these narrow laser beams
  - Distance from Earth to the Moon
  - Between continents to determine rate of plate movement
- Communications, Medical, Industrial, Surveying, Photography, Engineering
- A CD player reads small dot patterns that are converted into electronic signals, then sound

Heisenberg's Uncertainty Principle

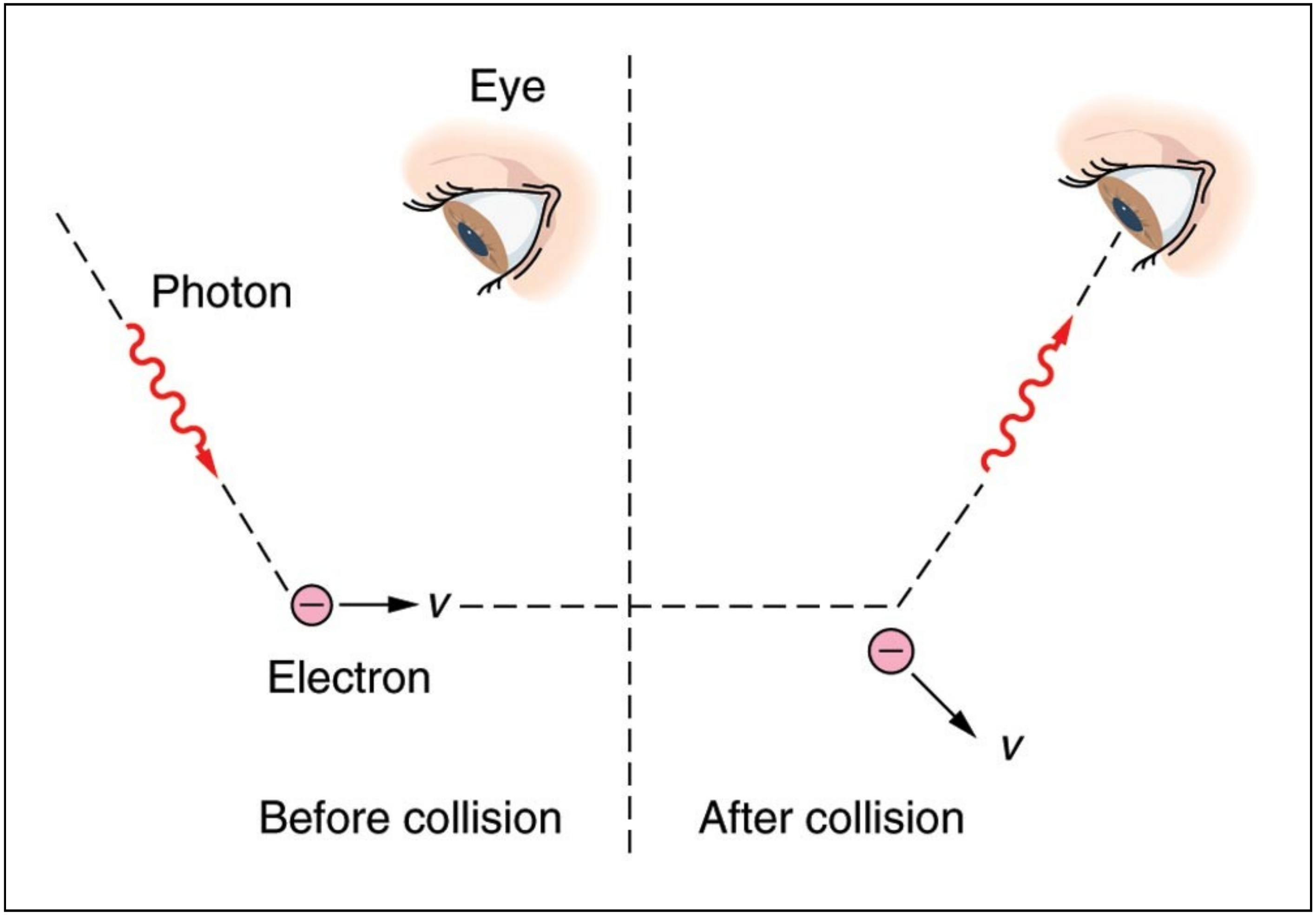
- In 1927 the German physicist introduced a new concept relating to measurement accuracy.
- Heisenberg's Uncertainty Principle can be stated as: It is impossible to know a particle's exact position and velocity simultaneously.

Audio Link

The very act of measurement may alter a particle's position and velocity.

- Suppose one is interested in the exact position and velocity of an electron.
- At least one photon must bounce off the electron and come to your eye.
- The collision process between the photon and the electron will alter the electron's position or velocity.

Bouncing a photon off the electron introduces a great deal of measurement uncertainty



How much does measurement alter the position and velocity?

- Further investigation led to the conclusion that several factors need to be considered in determining the accuracy of measurement:
  - Mass of the particle (m)
  - Minimum uncertainty in velocity (Δv)
  - Minimum uncertainty in position (Δx)
- When these three factors are multiplied together they equal a very small number.
  - Close to Planck's constant (h = 6.63 10^{-34} J·s)

Heisenberg's Uncertainty Principle

- Therefore: $m(\Delta v)(\Delta x) \cong h$
- Although this principle may be philosophically significant, it is only of practical importance when dealing with particles of atomic and subatomic size.

Matter Waves or de Broglie Waves

- With the development of the dual nature of light it became apparent that light “waves” sometime act like particles.
- Could the reverse be true?
- Can particles have wave properties?
- In 1925 the French physicist de Broglie postulated that matter has properties of both waves and particles.

De Broglie's Hypothesis

- Any moving particle has a wave associated with it whose wavelength is given by the following formula
- $\lambda = h/mv$
- λ = wavelength of the moving particle
- m = mass of the moving particle
- v = speed of the moving particle
- h = Planck's constant (6.63×10^{-34} J·s)

de Broglie Waves

- The waves associated with moving particles are called *matter waves* or *de Broglie waves*.
- Note in de Broglie's equation ($\lambda = h/mv$) the wavelength (λ) is inversely proportional to the mass of the particle (m)
- Therefore the longest wavelengths are associated with particles of very small mass.
- Also note that since h is so small, the resulting wavelengths of matter are also quite small.

Finding the de Broglie Wavelength

Exercise Example

- *Find the de Broglie wavelength for an electron (m = 9.11 x 10^{-31} kg) moving at 7.30 x 10^{5} m/s.*
- Use de Broglie equation: $\lambda = h/mv$
- We are given h, m, & v

$$\lambda = \frac{6.63 \times 10^{-34}\ \text{Js}}{(9.11 \times 10^{-31}\ \text{kg})(7.30 \times 10^{5}\ \text{m/s})}$$

Finding the de Broglie Wavelength
Exercise Example (cont.)

$\lambda = 1.0 \times 10^{-9}\text{m} = 1.0 \text{ nm (nanometer)}$

$$\lambda = \frac{6.63 \times 10^{-34} \text{ kg m}^2\text{s/s}^2}{(9.11 \times 10^{-31} \text{ kg})(7.30 \times 10^{5} \text{ m/s})}$$

This wavelength is only several times larger than the diameter of the average atom, therefore significant for an electron.

Finding the de Broglie Wavelength
Confidence Exercise

- *Find the de Broglie wavelength for a 1000 kg car traveling at 25 m/s*
- Use de Broglie equation: $\lambda = h/mv$
- We are given h, m, & v

$$\lambda = \frac{6.63 \text{ x } 10^{-34} \text{ Js}}{(1000 \text{ kg})(25 \text{ m/s})} = \frac{6.63 \text{ x } 10^{-34} \text{ kg m}^2\text{s/s}^2}{(1000 \text{ kg})(25 \text{ m/s})}$$

$$\lambda = 2.65 \text{ x } 10^{-38}\ m = 2.65 \text{ x } 10^{-29} \text{ nm}$$

A <u>very</u> short wavelength!

de Broglie's Hypothesis – Early Skepticism

- In 1927 two U.S. scientists, Davisson and Germer, experimentally verified that particles have wave characteristics.
- These two scientists showed that a bean of electrons (particles) exhibits a diffraction pattern (a wave property.)
- Recall Section 7.4 – appreciable diffraction only occurs when a wave passes through a slit of approximately the same width as the wavelength

de Broglie's Hypothesis – Verification

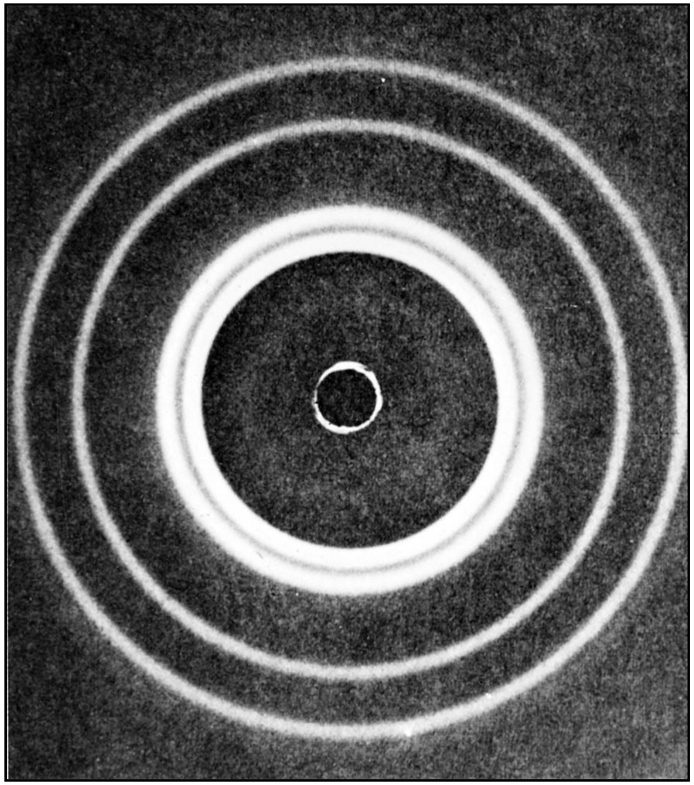
- Recall from our Exercise Example that an electron would be expected to have a $\lambda \cong 1$ nm.
- Slits in the range of 1 nm cannot be manufactured
- BUT … nature has already provided us with suitably small "slits" in the form of mineral crystal lattices.
- By definition the atoms in mineral crystals are arranged in an orderly and repeating pattern.

de Broglie's Hypothesis – Verification

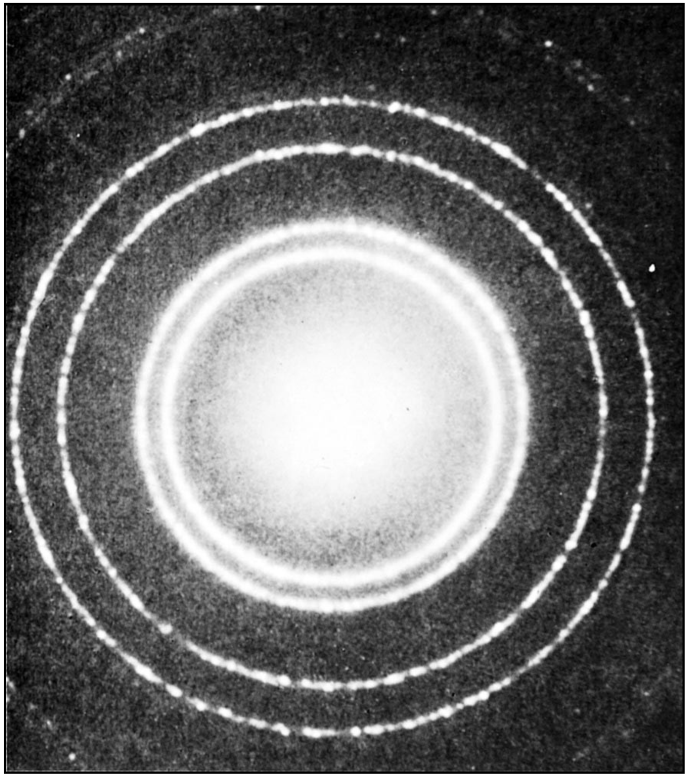
- The orderly rows within a crystal lattice provided the extremely small slits needed (in the range of 1 nm.)
- Davisson and Germer photographed two diffraction patterns.
  - One pattern was made with X-rays (waves) and one with electrons (particles.)
- The two diffraction patterns are remarkably similar.

Similar Diffraction Patterns
Both patterns indicate wave-like properties

X-Ray pattern



Diffraction pattern of electrons



Dual Nature of Matter

- Electron diffraction demonstrates that moving matter not only has particle characteristics, but also wave characteristics
- BUT …
- The wave nature of matter only becomes of practical importance with extremely small particles such as electrons and atoms.

Electron Microscope

- The electron microscope is based on the principle of matter waves.
- This device uses a beam of electrons to view objects.
- Recall that the wavelength of an electron is in the order of 1 nm, whereas the wavelength of visible light ranges from 400 – 700 nm

Electron Microscope

- The amount of fuzziness of an image is directly proportional to the wavelength used to view it
- therefore …
- the electron microscope is capable of much finer detail and greater magnification than a microscope using visible light.

Electron Cloud Model of an Atom

- Recall that Bohr chose to analyze the hydrogen atom, because it is the simplest atom
- It is increasingly difficult to analyze atoms with more that one electron, due to the myriad of possible electrical interactions
- In large atoms, the electrons in the outer orbits are also partially shielded from the attractive forces of the nucleus

Audio Link

Electron Cloud Model of an Atom

- Although Bohr's theory was very successful in explaining the hydrogen atom …
- This same theory did not give correct results when applied to multielectron atoms
- Bohr was also unable to explain *why* the electron energy levels were quantized
- Additionally, Bohr was unable to explain *why* the electron did not radiate energy as it traveled in its orbit

Bohr's Theory – Better Model Needed

- With the discovery of the dual natures of both waves and particles …
- A new kind of physics was developed, called quantum mechanics or wave mechanics
  - Developed in the 1920's and 1930's as a synthesis of wave and quantum ideas
- Quantum mechanisms also integrated Heisenberg's uncertainty principle
  - The concept of probability replaced the views of classical mechanics in describing electron movement

Quantum Mechanics

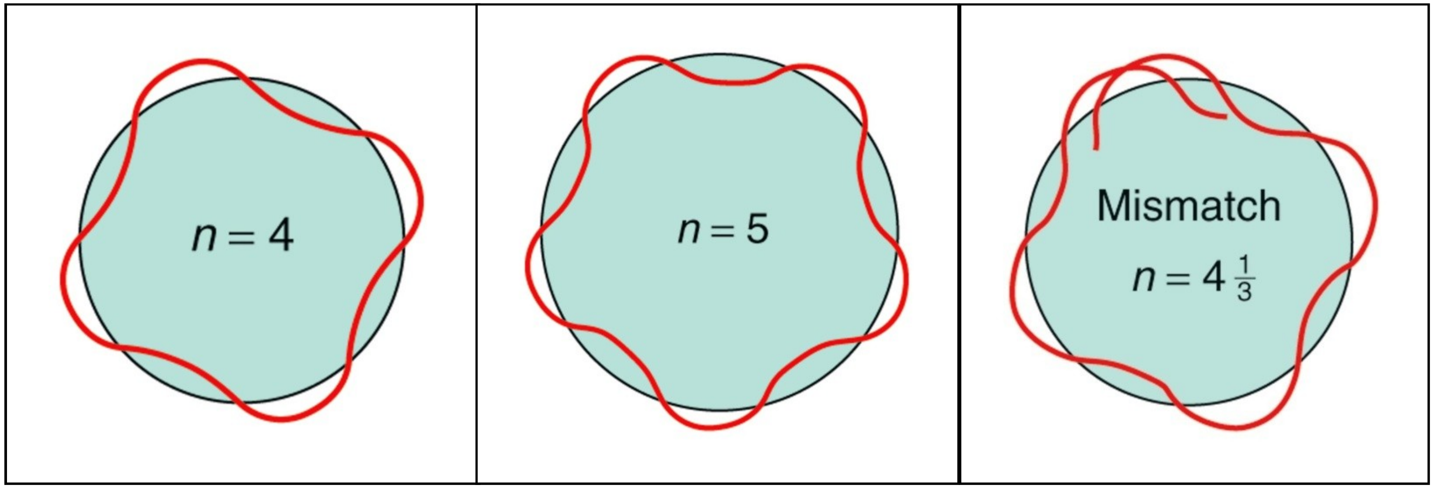
- In 1926, the Austrian physicist Erwin Schrödinger presented a new mathematical equation applying de Broglie’s matter waves.
- Schrödinger’s equation was basically a formulation of the conservation of energy
- The simplified form of this equation is …
- $(E_k + E_p)\Psi = E\Psi$
  - E_k, E_p, and E are kinetic, potential, and total energies, respectively
  - Ψ = wave function

Quantum Mechanical Model *or Electron Cloud Model*

- Schrödinger’s model focuses on the wave nature of the electron and treats it as a standing wave in a circular orbit
- Permissible orbits must have a circumference that will accommodate a whole number of electron wavelengths (λ)
- If the circumference will not accommodate a whole number λ, then this orbit is not ’probable’

The Electron as a Standing Wave

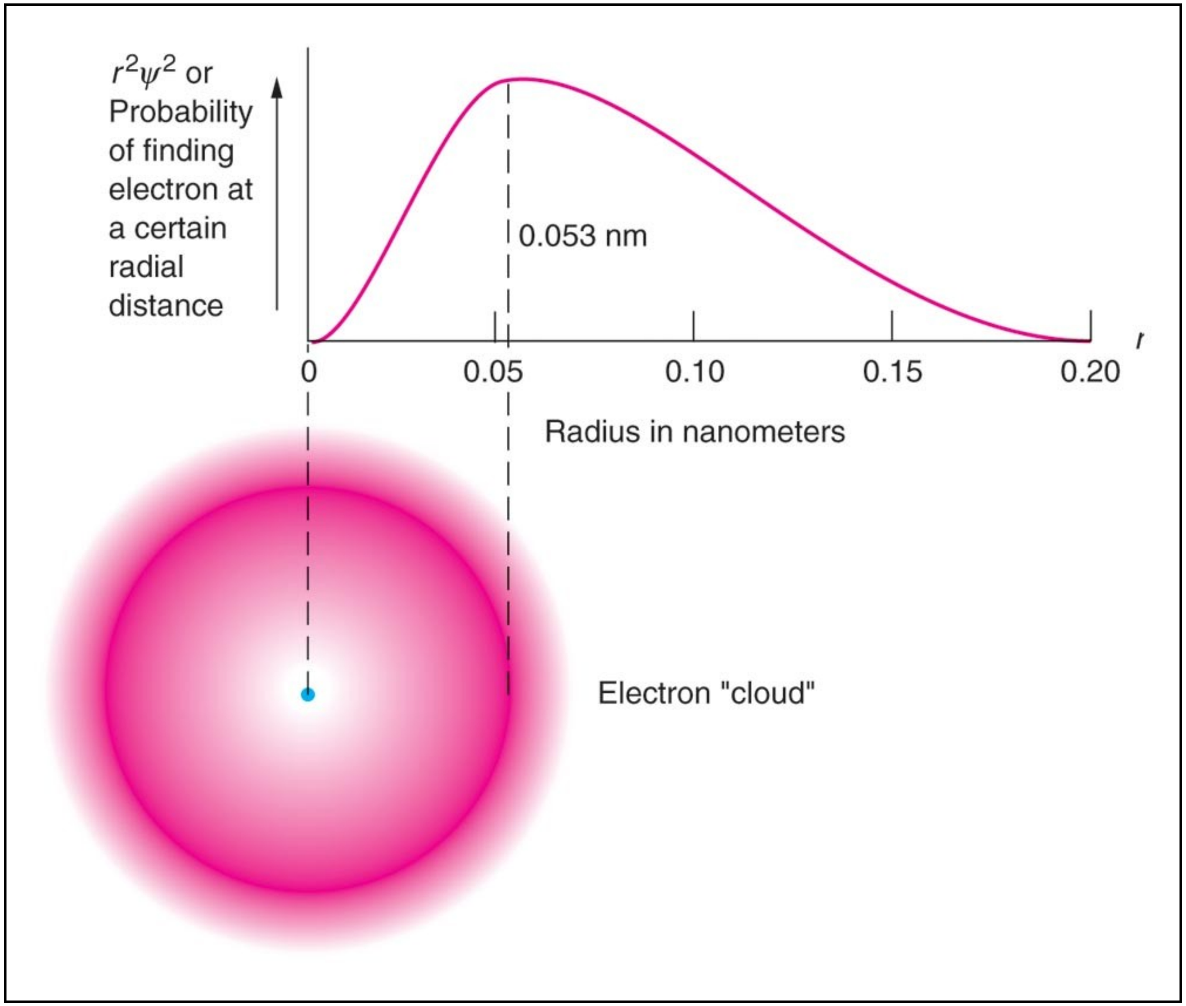
- For the electron wave to be stable, the circumference of the orbit must be a whole number of wavelengths



Wave Function & Probability

- Mathematically, the wave function (Ψ) represents the wave associated with a particle
- For the hydrogen atom it was found that the equation $r^2\Psi^2$ represents …
  - The probability of the hydrogen electron being a certain distance *r* from the nucleus
- A plot of $r^2\Psi^2$ versus *r* for the hydrogen electron shows that the most probable radius for the hydrogen electron is $r = 0.053$nm
  - Same value as Bohr predicted in 1913!

$r^2\Psi^2$ (Probability) versus r (*Radius*)



Concept of the *Electron Cloud*

- Although the hydrogen electron may be found at a radii other than 0.053 nm – the probability is lower
- Therefore, when viewed from a probability standpoint, the “electron cloud” around the nucleus represents the probability that the electron will be at that position
- The electron cloud is actually a visual representation of a probability distribution

Changing Model of the Atom

- Although Bohr’s “planetary model” was brilliant and quite elegant it was not accurate for multielectron atoms
- Schrödinger’s model is highly mathematical and takes into account the electron’s wave nature

Schrödinger's Quantum Mechanical Model

- The quantum mechanical model only gives the location of the electrons in terms of probability
- But, this model enables scientists to determine accurately the energy of the electrons in multielectron atoms
- Knowing the electron's energy is much more important than knowing its precise location

Chapter 9 - Important Equations

- $E = hf$ Photon Energy (h= 6.63 x 10^{-34} Js)
- $r_n = 0.053\ n^2$ nm Hydrogen Electron Orbit Radii
- E_n = (-13.60/n2) eV Hydrogen Electron Energy
- $E_{\text{photon}} = E_{n_i} - E_{n_f}$ Photon Energy for Transition
- $\lambda = h/mv$ de Broglie Wavelength