Lecture 10

Atomic and Quantum Physics
Ch 27 & Ch 28

You may ignore 28.6 and 28.7

Need for Quantum Physics

- Problems remained from classical mechanics that relativity didn't explain.
- Blackbody Radiation
  - The electromagnetic radiation emitted by a heated object
- Photoelectric Effect
  - Emission of electrons by an illuminated metal
- Spectral Lines
  - Emission of sharp spectral lines by gas atoms in an electric discharge tube

Development of Quantum Physics

- 1900 to 1930
  - Development of ideas of quantum mechanics
    - Also called wave mechanics
    - Highly successful in explaining the behavior of atoms, molecules, and nuclei
- Involved a large number of physicists
  - Planck introduced basic ideas.
  - Mathematical developments and interpretations involved such people as Einstein, Bohr, Schrödinger, de Broglie, Heisenberg, Born and Dirac.

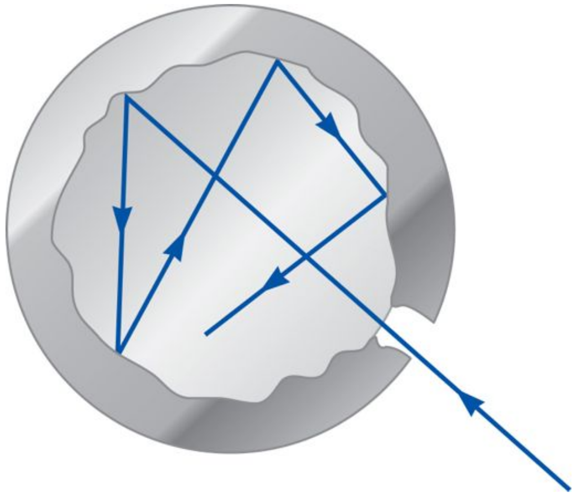
Blackbody Radiation

Blackbody Radiation

- An object at any temperature emits electromagnetic radiation.
  - Also called *thermal radiation.*
  - Stefan's Law describes the total power radiated.
  - The spectrum of the radiation depends on the temperature and properties of the object.
    - The spectrum shows a continuous distribution of wavelengths from infrared to ultaviolet.

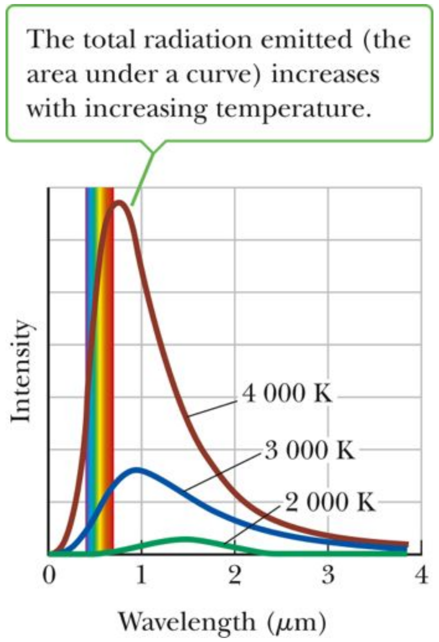
Blackbody Radiation – Classical View

- Thermal radiation originates from accelerated charged particles.
- Problem in explaining the observed energy distribution
- Opening in a cavity is a good approximation
- The nature of the radiation emitted through the opening depends only on the temperature of the cavity walls.



Blackbody Radiation Graph

- Experimental data for distribution of energy in blackbody radiation
- As the temperature increases, the total amount of energy increases.
  - Shown by the area under the curve
- As the temperature increases, the peak of the distribution shifts to shorter wavelengths.

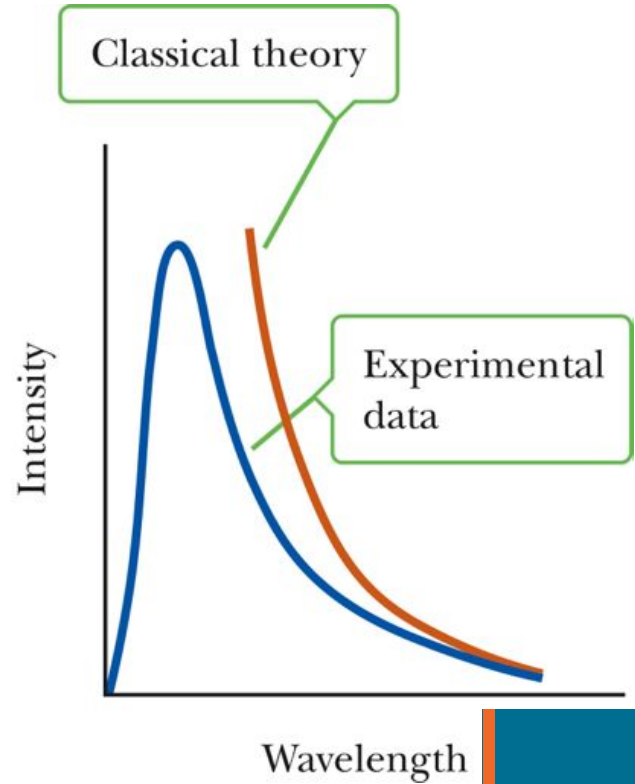


Wien's Displacement Law

- The wavelength of the peak of the blackbody distribution was found to follow *Wein's Displacement Law.*
  - $\lambda_{max} T = 0.2898 \times 10^{-2}$ m • K
    - λ_{max} is the wavelength at which the curve peaks.
    - T is the absolute temperature of the object emitting the radiation.

The Ultraviolet Catastrophe

- Classical theory did not match the experimental data.
- At long wavelengths, the match is good.
- At short wavelengths, classical theory predicted infinite energy.
- At short wavelengths, experiment showed no energy

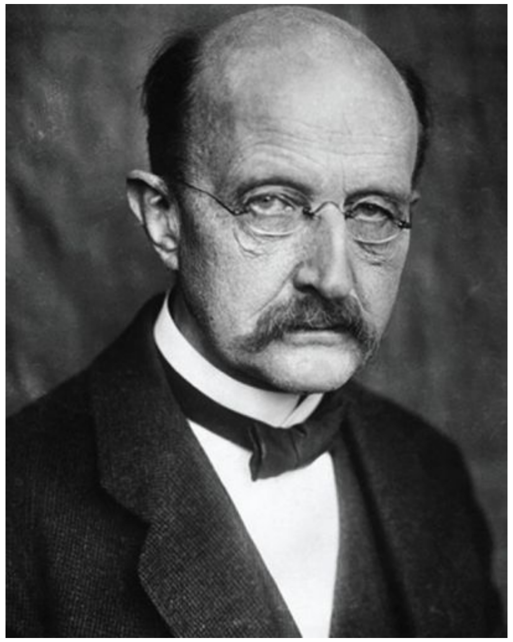


Planck's Resolution

- Planck hypothesized that the blackbody radiation was produced by *resonators.*
  - Resonators were submicroscopic charged oscillators.
- The resonators could only have *discrete energies.*
  - $E_n = n\,h\,f$
    - n is called the *quantum number*
    - ƒ is the frequency of vibration
    - h is *Planck's constant*, 6.626×10^{-34} J s
- Key point is quantized energy states

Max Planck

- 1858 – 1947
- Introduced a “quantum of action,” h
- Awarded Nobel Prize in 1918 for discovering the quantized nature of energy



Quantized Energy

- Planck’s assumption of quantized energy states was a radical departure from classical mechanics.
- The fact that energy can assume only certain, discrete values is the single most important difference between quantum and classical theories.
  - Classically, the energy can be in any one of a continuum of values.

Hw 4 & 5

If a quantum of radiation has an energy of 2.0 keV, what is its wavelength? (h = 6.63 x 10⊠^{34} Js, 1 eV = 1.60 x 10-^{19} J, c = 3.00 x10^{8} m/s, and 1 nm = 10-^{9} m)
a.0.32 nm
b.0.41 nm
c.0.62 nm
d.1.02 nm

What is the surface temperature of a distant star (which emits light as if it were a blackbody) where the peak wavelength is 480 nm? (Hint: The surface of the human body at 35° C has a peak wavelength of 941 ìm). (1 nm = 10-^{9} m = 10^{-3} ìm)
a.4 510 K
b.5 100 K
c.6 040 K
d.6 350 K

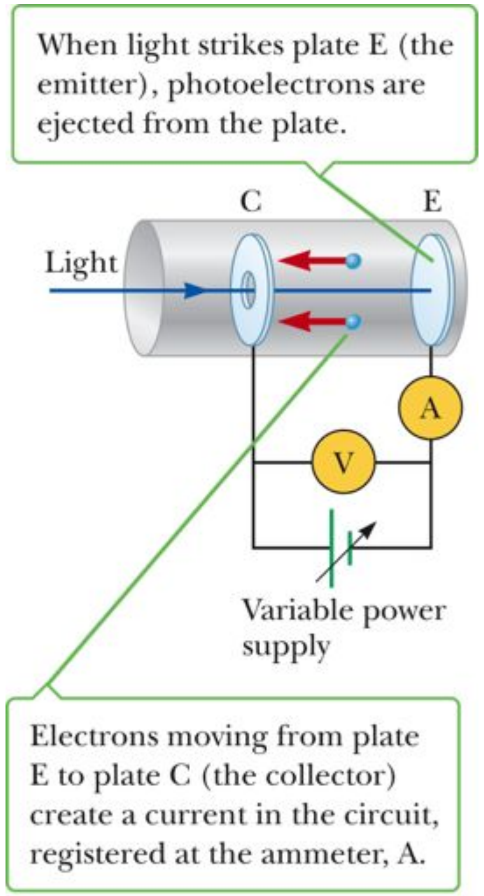
Star A has the peak of its blackbody radiation at ë$_{A}$. Star B has its peak at ë$_{B}$, which is one-fourth that of ë$_{A}$. If Star A's surface temperature is T_A, how does the surface temperature T_B of Star B compare?
a.$T_B = 16\ T_A$
b.$T_B = 4\ T_A$
c.$T_B = T_A/4$
d.$T_B = T_A/16$

A quantum of radiation has an energy of 2.0 keV. What is its frequency? (h = 6.63 x 10-^{34} Js and 1 eV = 1.60 x 10-^{19} J)
a.3.2 x 10^{17} Hz
b.4.8 x 10^{17} Hz
c.6.3 x 10^{17} Hz
d.7.3 x 10^{17} Hz

Photoelectric Effect

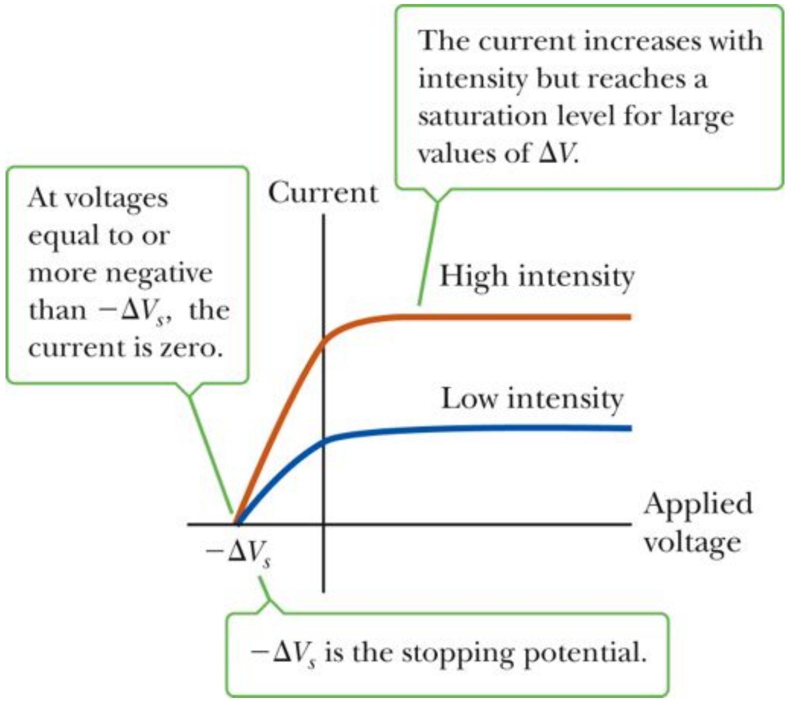
Photoelectric Effect Schematic

- When light strikes E, photoelectrons are emitted.
- Electrons collected at C and passing through the ammeter create a current in the circuit.
- C is maintained at a positive potential by the power supply.



Photoelectric Current/Voltage Graph

- The current increases with intensity, but reaches a saturation level for large ΔV' s.
- No current flows for voltages less than or equal to $-\Delta V_s$, the *stopping potential.*



More About Photoelectric Effect

- The stopping potential is independent of the radiation intensity.
- The maximum kinetic energy of the photoelectrons is related to the stopping potential: $KE_{max} = e\Delta V_s$

Features Not Explained by Classical Physics/Wave Theory

- No electrons are emitted if the incident light frequency is below some *cutoff frequency* that is characteristic of the material being illuminated.
- The maximum kinetic energy of the photoelectrons is independent of the light intensity.

Intensity vs wavelength?

More Features Not Explained

- The maximum kinetic energy of the photoelectrons increases with increasing light frequency.
- Electrons are emitted from the surface almost instantaneously, even at low intensities.

Einstein's Explanation

- A tiny packet of light energy, called a photon, would be emitted when a quantized oscillator jumped from one energy level to the next lower one.
  - Extended Planck's idea of quantization to electromagnetic radiation
- The photon's energy would be $E = hf$
- Each photon can give all its energy to an electron in the metal.
- The maximum kinetic energy of the liberated photoelectron is $KE_{max} = hf - \phi$
- ϕ is called the *work function* of the metal

Explanation of Classical "Problems"

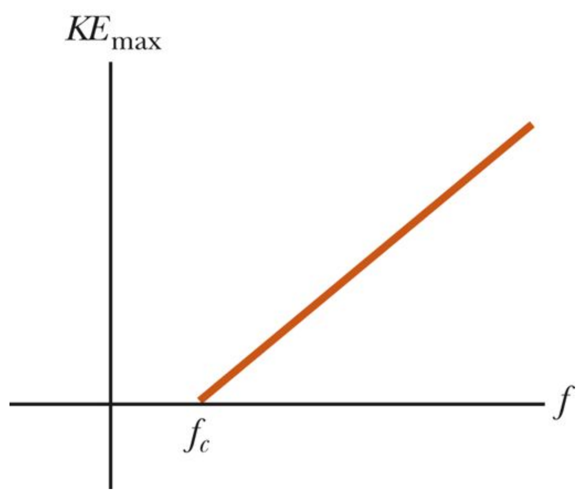
- The effect is not observed below a certain cutoff frequency since the photon energy must be greater than or equal to the work function.
  - Without this, electrons are not emitted, regardless of the intensity of the light
- The maximum KE depends only on the frequency and the work function, not on the intensity.
  - The absorption of a single photon is responsible for the electron's kinetic energy.

More Explanations

- The maximum KE increases with increasing frequency.
- The effect is instantaneous since there is a one-to-one interaction between the photon and the electron.

Verification of Einstein's Theory

- Experimental observations of a linear relationship between KE and frequency confirm Einstein's theory.
- The x-intercept is the cutoff frequency.



Cutoff Wavelength

- The cutoff wavelength is related to the work function.

$$\lambda_c = \frac{hc}{\phi}$$

- Wavelengths greater than λ_C incident on a material with a work function ϕ don't result in the emission of photoelectrons.

Photocells

- Photocells are an application of the photoelectric effect.
- When light of sufficiently high frequency falls on the cell, a current is produced.
- Examples
  - Streetlights, garage door openers, elevators

Hw 9 & 10

If a monochromatic light beam with quantum energy value of 3.0 eV incident upon a photocell where the work function of the target metal is 1.60 eV, what is the maximum kinetic energy of ejected electrons?
a.4.6 eV
b.4.8 eV
c.1.4 eV
d.2.4 eV

According to Einstein, as the wavelength of the incident monochromatic light beam becomes shorter, the work function of a target material in a phototube:
a.increases.
b.decreases.
c.remains constant.
d.is directly proportional to wavelength.

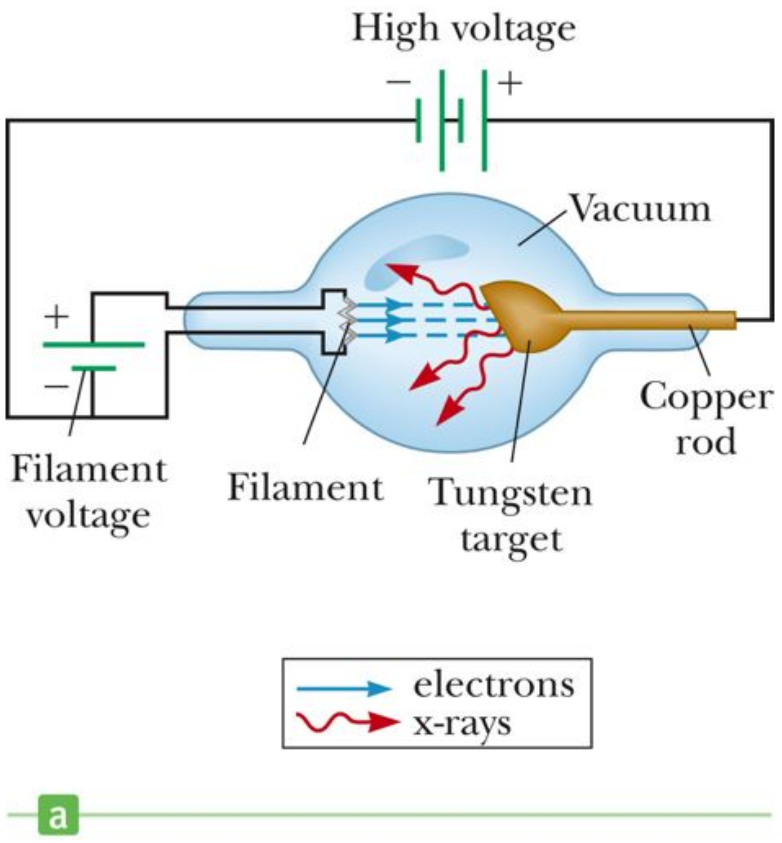
A monochromatic light beam is incident on a barium target, which has a work function of 2.50 eV. If a stopping potential of 1.0 V is required, what is the light beam photon energy?
a.1.0 eV
b.1.5 eV
c.2.5 eV
d.3.5 eV

X-Rays

- Discovered and named by RÖntgen in 1895
- Later identified as electromagnetic radiation with short wavelengths
  - Wavelengths lower (frequencies higher) than for ultraviolet
  - Wavelengths are typically about 0.1 nm.
  - X-rays have the ability to penetrate most materials with relative ease.

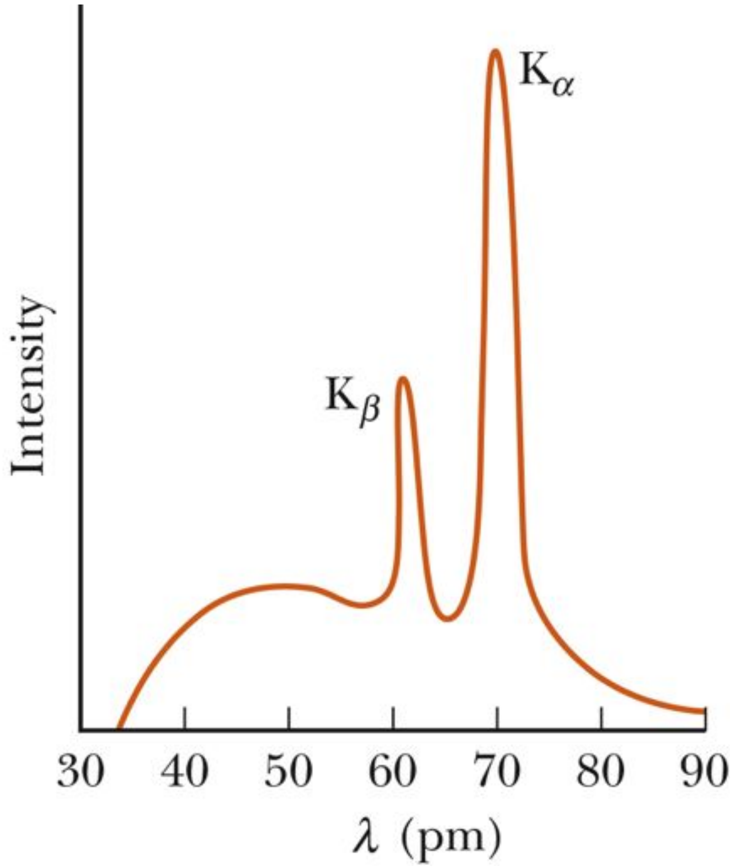
Production of X-rays, 1

- X-rays are produced when high-speed electrons are suddenly slowed down.
  - Can be caused by the electron striking a metal target
- Heat generated by current in the filament causes electrons to be emitted.
- These freed electrons are accelerated toward a dense metal target.
- The target is held at a higher potential than the filament.



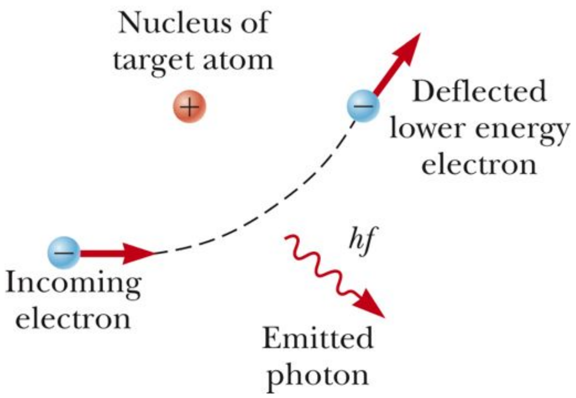
X-ray Spectrum

- The x-ray spectrum has two distinct components.
- Continuous broad spectrum
  - Depends on voltage applied to the tube
  - Sometimes called **bremsstrahlung**
- The sharp, intense lines depend on the nature of the target material.



Production of X-rays, 2

- An electron passes near a target nucleus.
- The electron is deflected from its path by its attraction to the nucleus.
  - This produces an acceleration
- It will emit electromagnetic radiation when it is accelerated.



Wavelengths Produced

- If the electron loses all of its energy in the collision, the initial energy of the electron is completely transformed into a photon.
- The wavelength can be found from

$$e\Delta V = hf_{\max} = \frac{hc}{\lambda_{\min}}$$

Wavelengths Produced, Cont.

- Not all radiation produced is at this minimum wavelength.
  - Many electrons undergo more than one collision before being stopped.
  - This results in the continuous spectrum produced.

X-Ray Diffraction

Diffraction of X-rays by Crystals

- For diffraction to occur, the spacing between the lines must be approximately equal to the wavelength of the radiation to be measured.
- The regular array of atoms in a crystal can act as a three-dimensional grating for diffracting X-rays.

Schematic for X-ray Diffraction

- A beam of X-rays with a continuous range of wavelengths is incident on the crystal.
- The diffracted radiation is very intense in certain directions.
  - These directions correspond to constructive interference from waves reflected from the layers of the crystal.
- The diffraction pattern is detected by photographic film.

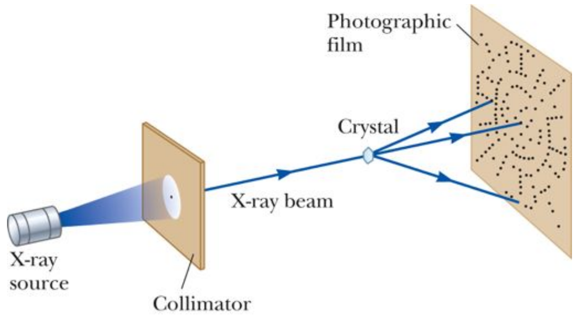
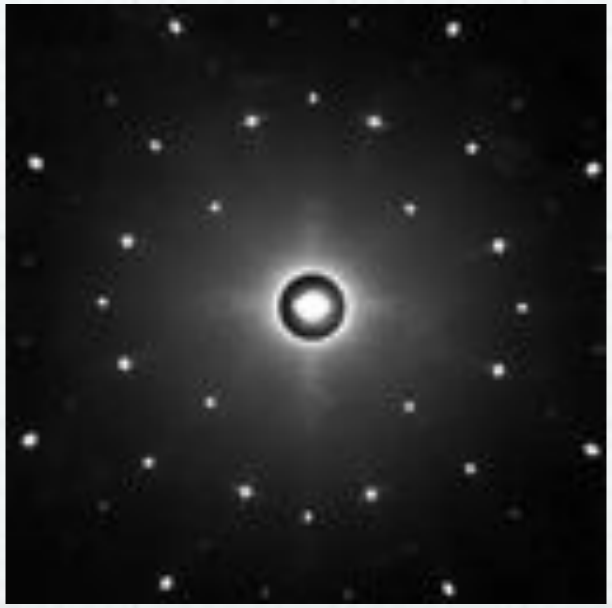


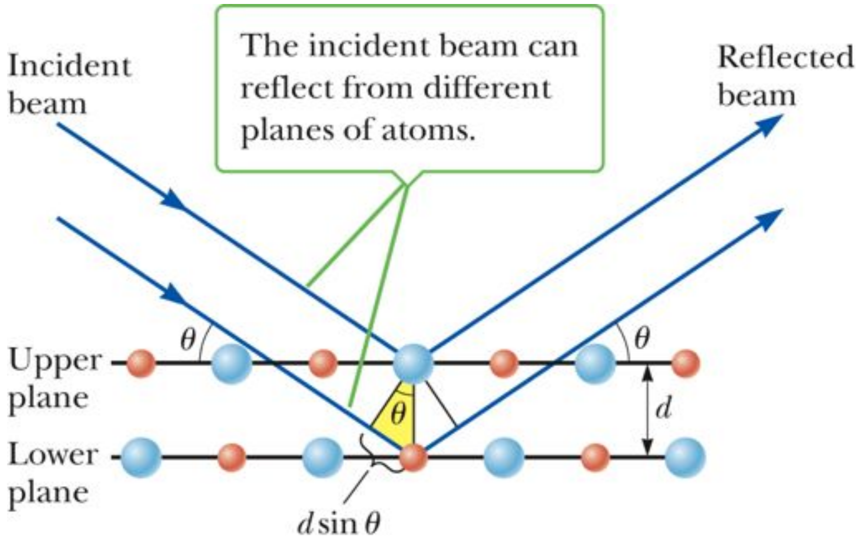
Photo of X-ray Diffraction Pattern

- The array of spots is called a *Laue* pattern.
- The crystal structure is determined by analyzing the positions and intensities of the various spots.



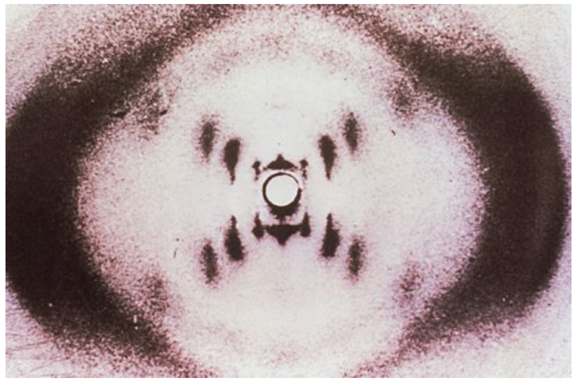
Bragg’s Law

- The beam reflected from the lower surface travels farther than the one reflected from the upper surface.
- If the path difference equals some integral multiple of the wavelength, constructive interference occurs.
- *Bragg’s Law* gives the conditions for constructive interference.
  - $2d \sin \theta = m\lambda$, $m = 1, 2, 3\ldots$



Uses of X-Ray Diffraction

- X-ray diffraction is used to determine the molecular structure of proteins, DNA, and RNA.
- X-rays with $\lambda = 0.10$ nm are used.
- The geometry of the diffraction pattern is determined by the lattice arrangement of the molecules.
- The intensities are determined by the atoms and their electronic distribution in the cell.
- Picture shows an x-ray diffraction photo of DNA



Hw 17 & 19

The spacing between atoms in KCl crystal is 3.1×10^{-10} m. At what angle from the surface will a beam of 3.14×10^{-11} m x-rays be constructively scattered?

a. 57°

b. 2.9°

c. 90°

d. 10°

What is the highest frequency of the photons produced by a 90-kV x-ray machine? ($h = 6.63 \times 10^{-34}$ Js)

a. 1.2×10^{19} Hz

b. 1.1×10^{19} Hz

c. 2.4×10^{19} Hz

d. 2.2×10^{19} Hz

Compton Effect

Arthur Holly Compton

- 1892 – 1962
- Discovered the Compton effect
- Worked with cosmic rays
- Director of the lab at U of Chicago
- Shared Nobel Prize in 1927

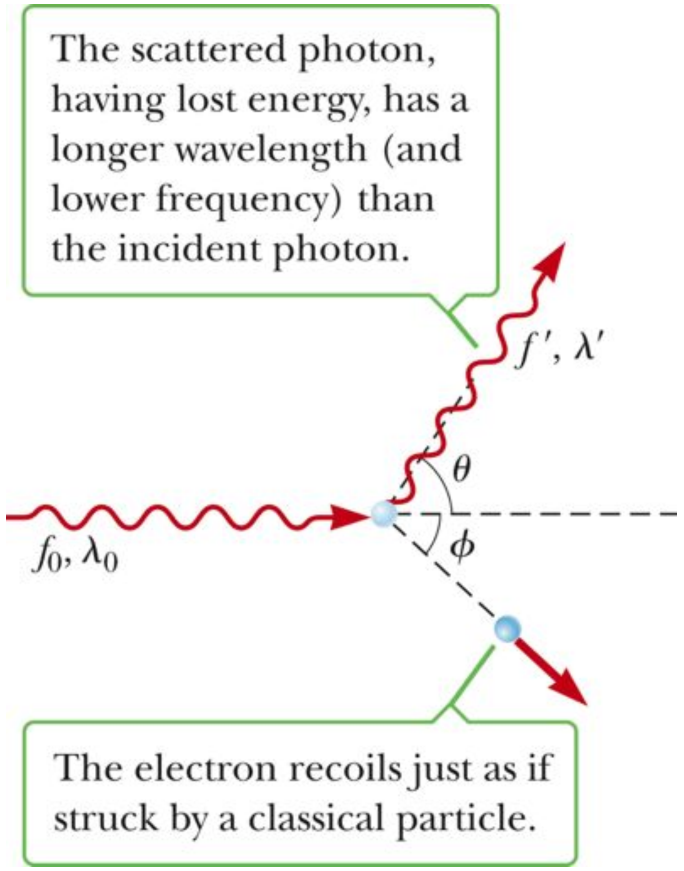
The Compton Effect

- Compton directed a beam of x-rays toward a block of graphite.
- He found that the scattered x-rays had a slightly longer wavelength that the incident x-rays.
  - This means they also had less energy.
- The amount of energy reduction depended on the angle at which the x-rays were scattered.
- The change in wavelength is called the *Compton shift*.

Compton Scattering

- Compton assumed the photons acted like other particles in collisions.
- Energy and momentum were conserved.
- The shift in wavelength is

$$\Delta\lambda = \lambda - \lambda_o = \frac{h}{m_e c}(1-\cos\theta)$$



Compton Scattering, Final

- The quantity h/m_ec is called the *Compton wavelength.*
  - Compton wavelength = 0.002 43 nm
  - Very small compared to visible light
- The Compton shift depends on the scattering angle and not on the wavelength.
- Experiments confirm the results of Compton scattering and strongly support the photon concept.

Hw 23

In the Compton effect, what is the greatest change in wavelength that can occur? ($h = 6.63 \times 10^{-34}$ J·s, $m_{electron} = 9.11 \times 10^{-31}$ kg, and $c = 3.00 \times 10^{8}$ m/s)

a. 2.43×10^{-12} m

b. 4.85×10^{-12} m

c. equal to the incident wavelength

d. infinite

The dual nature of light and matter

Photons and Electromagnetic Waves

- **Light has a dual nature. It exhibits both wave and particle characteristics.**
  - Applies to all electromagnetic radiation
  - Different frequencies allow one or the other characteristic to be more easily observed.
- The photoelectric effect and Compton scattering offer evidence for the particle nature of light.
  - When light and matter interact, light behaves as if it were composed of particles.
- Interference and diffraction offer evidence of the wave nature of light.

Louis de Broglie

- 1892 – 1987
- Discovered the wave nature of electrons
- Awarded Nobel Prize in 1929

Wave Properties of Particles

- In 1924, Louis de Broglie postulated that **because photons have wave and particle characteristics, perhaps all forms of matter have both properties.**
- Furthermore, the frequency and wavelength of matter waves can be determined.

de Broglie Wavelength and Frequency

- The *de Broglie wavelength* of a particle is

$$\lambda = \frac{h}{p} = \frac{h}{mv}$$

- The frequency of matter waves is

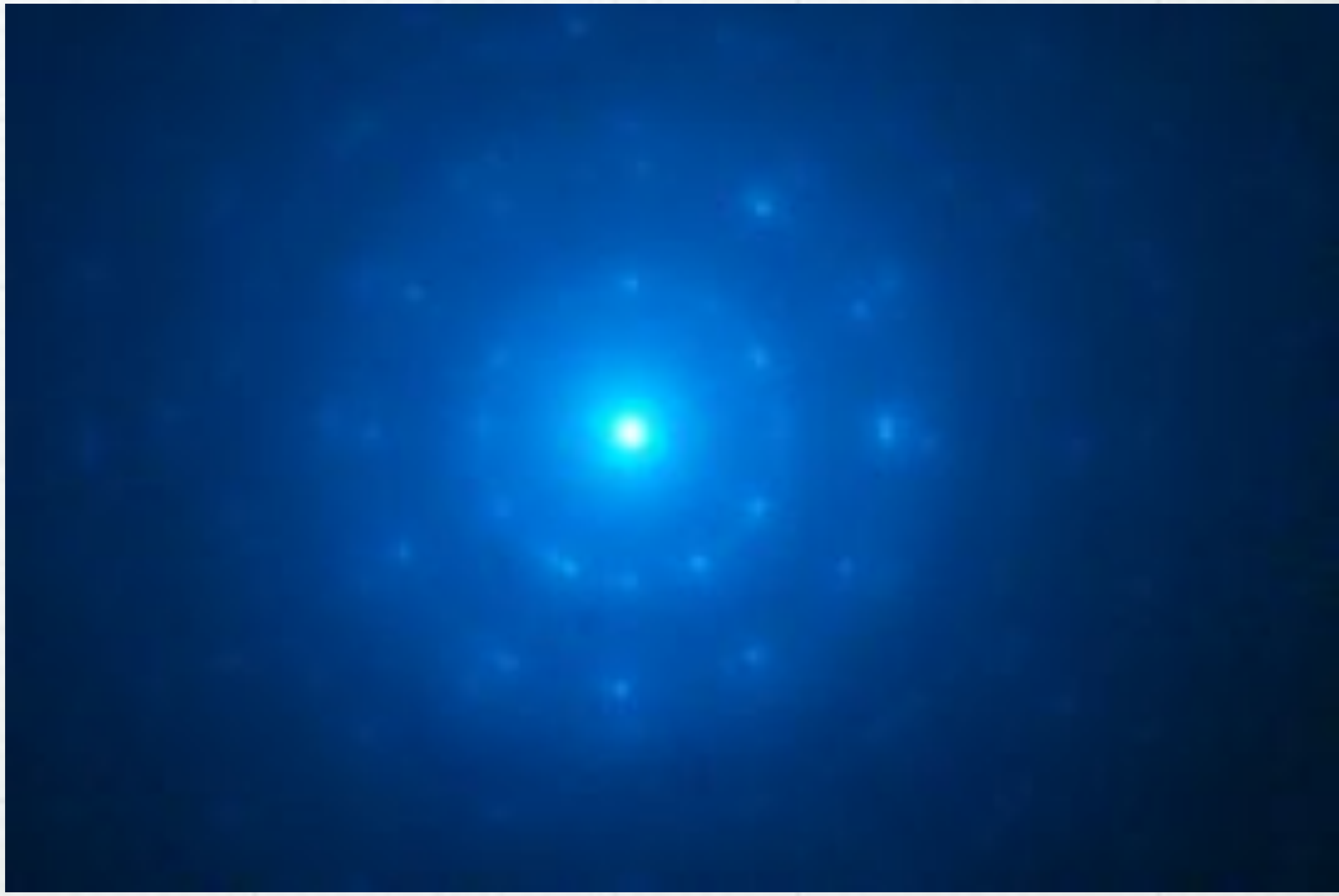
$$f = \frac{E}{h}$$

Dual Nature of Matter

- The de Broglie equations show the dual nature of matter.
- Each contains matter concepts.
  - Energy and momentum
- Each contains wave concepts.
  - Wavelength and frequency

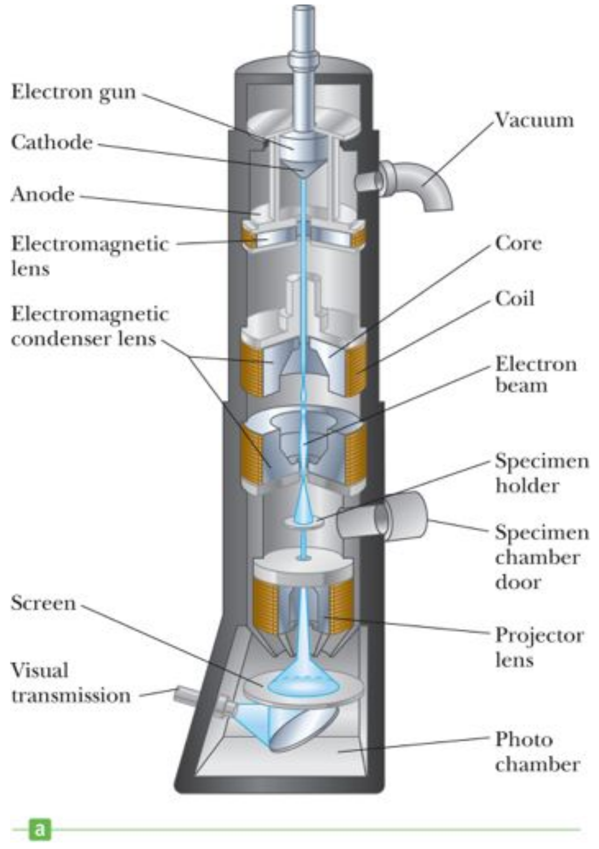
The Davisson-Germer Experiment

- They scattered low-energy electrons from a nickel target.
- They followed this with extensive diffraction measurements from various materials.
- The wavelength of the electrons calculated from the diffraction data agreed with the expected de Broglie wavelength.
- This confirmed the wave nature of electrons.



The Electron Microscope

- The electron microscope depends on the wave characteristics of electrons.
- Microscopes can only resolve details that are slightly smaller than the wavelength of the radiation used to illuminate the object.
- The electrons can be accelerated to high energies and have small wavelengths.



Introduction to Quantum Mechanics

Erwin Schrödinger

- 1887 – 1961
- Best known as the creator of wave mechanics
- Worked on problems in general relativity, cosmology, and the application of quantum mechanics to biology

Section 27.7

The Wave Function

- In 1926 Schrödinger proposed a wave equation that describes the manner in which matter waves change in space and time.
- Schrödinger's wave equation is a key element in quantum mechanics.
- Schrödinger's wave equation is generally solved for the *wave function*, Ψ.

The Wave Function, Cont.

- The wave function depends on the particle's position and the time.
- The value of Ψ^2 at some location at a given time is proportional to the probability of finding the particle at that location at that time.
  - Actually gives the probability per unit volume

The Uncertainty Principle

- When measurements are made, the experimenter is always faced with experimental uncertainties in the measurements.
  - Classical mechanics offers no fundamental barrier to ultimate refinements in measurements.
  - Classical mechanics would allow for measurements with arbitrarily small uncertainties.

The Uncertainty Principle, 2

- Quantum mechanics predicts that a barrier to measurements with ultimately small uncertainties does exist.
- In 1927 Heisenberg introduced the *uncertainty principle.*
  - If a measurement of position of a particle is made with precision Δx and a simultaneous measurement of linear momentum is made with precision Δp_x, then the product of the two uncertainties can never be smaller than h/4π

The Uncertainty Principle, 3

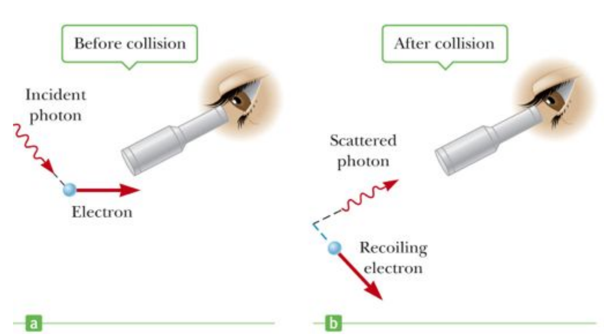
- Mathematically,

$$\Delta x \Delta p_x \geq \frac{h}{4\pi}$$

  - It is physically impossible to measure simultaneously the exact position and the exact linear momentum of a particle.
- Another form of the principle deals with energy and time:

$$\Delta E \Delta t \geq \frac{h}{4\pi}$$

Thought Experiment – The Uncertainty Principle



- A thought experiment for viewing an electron with a powerful microscope
- In order to see the electron, at least one photon must bounce off it.
- During this interaction, momentum is transferred from the photon to the electron.
- Therefore, the light that allows you to accurately locate the electron changes the momentum of the electron.

Uncertainty Principle Applied to an Electron

- View the electron as a particle.
- Its position and velocity cannot both be known precisely at the same time.
- Its energy can be uncertain for a period given by $\Delta t = h / (4 \pi \Delta E)$

Hw 29 & 31

What is the de Broglie wavelength for a proton ($m = 1.67 \times 10^{-27}$ kg) moving at a speed of 6.0 $\times 10^{6}$ m/s?
($h = 6.63 \times 10^{-34}$ Js)
a. 2.0×10^{-13} m
b. 0.33×10^{-13} m
c. 1.3×10^{-13} m
d. 0.66×10^{-13} m

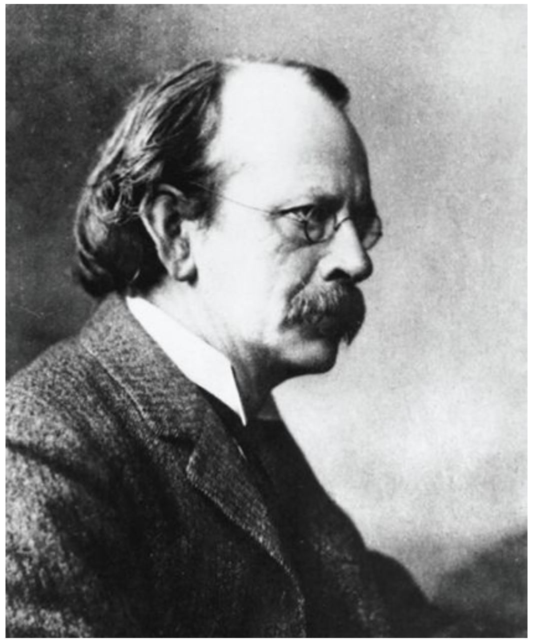
If an electron has a measured wavelength of 0.850×10^{10} m, what is its kinetic energy?
($h = 6.63 \times 10^{-34}$ Js, 1 eV = 1.6×10^{-19} J, and $m_e = 9.11 \times 10^{-31}$ kg)
a. 55.0 eV
b. 104 eV
c. 147 eV
d. 209 eV

Quantum Numbers and Atomic Structure

- The characteristic wavelengths emitted by a hot gas can be understood using quantum numbers.
- No two electrons can have the same set of quantum numbers – helps us understand the arrangement of the periodic table.
- Atomic structure can be used to describe the production of x-rays and the operation of a laser.

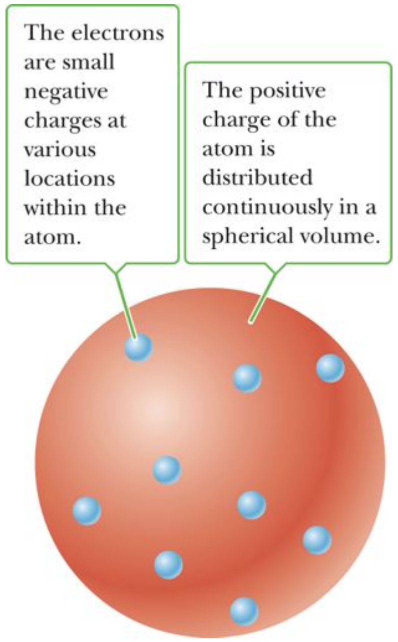
Sir Joseph John Thomson

- "J. J." Thomson
- 1856 - 1940
- Discovered the electron
- Did extensive work with cathode ray deflections
- 1906 Nobel Prize for discovery of electron

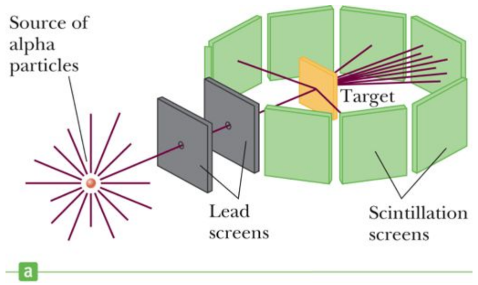


Early Models of the Atom

- J.J. Thomson's model of the atom
  - A volume of positive charge
  - Electrons embedded throughout the volume
- A change from Newton's model of the atom as a tiny, hard, indestructible sphere



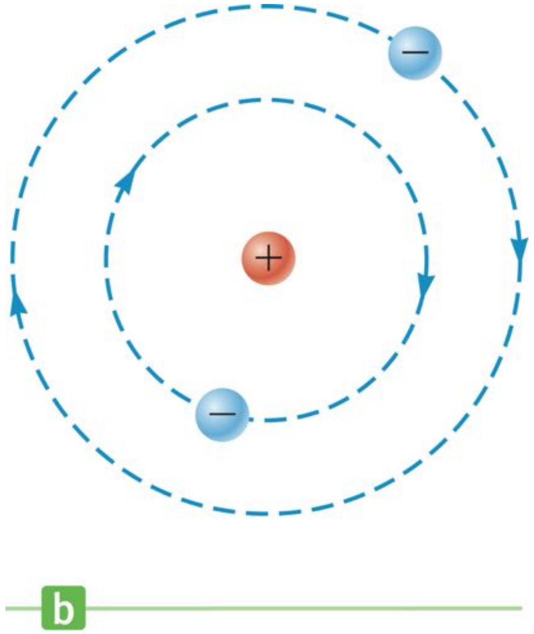
Scattering Experiments



- The source was a naturally radioactive material that produced alpha particles.
- Most of the alpha particles passed though the foil.
- A few deflected from their original paths
  - Some even reversed their direction of travel.

Early Models of the Atom, 2

- Rutherford, 1911
  - Planetary model
  - Based on results of thin foil experiments
  - Positive charge is concentrated in the center of the atom, called the *nucleus.*
  - Electrons orbit the nucleus like planets the sun.



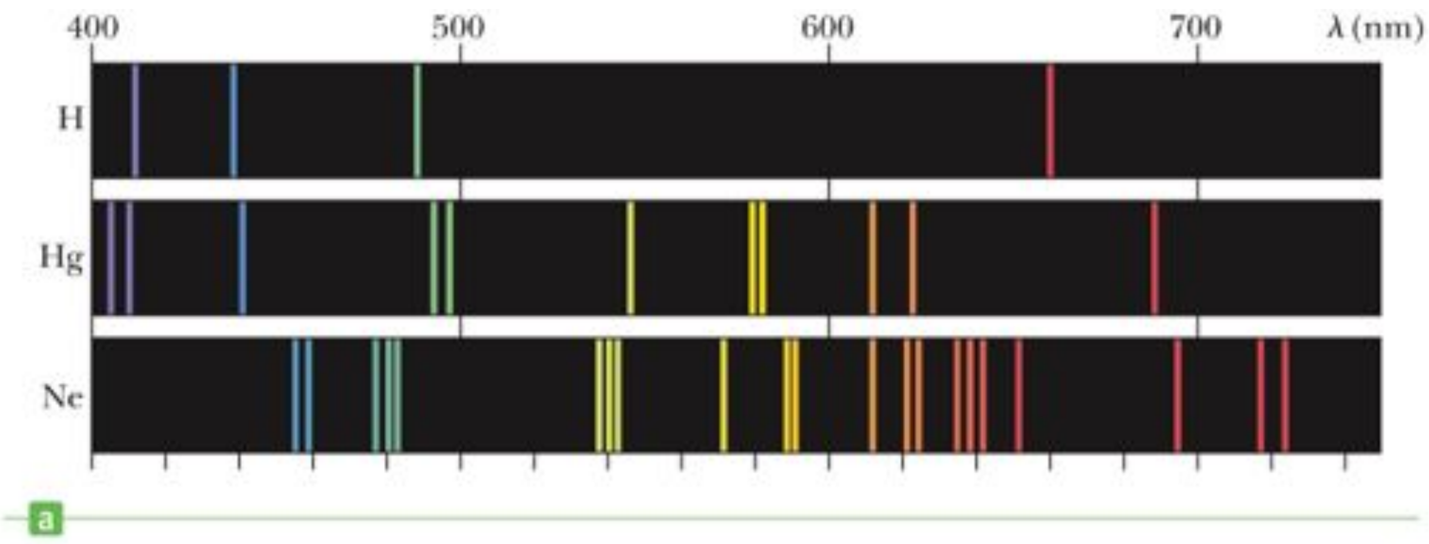
Difficulties with the Rutherford Model

- Atoms emit certain discrete characteristic frequencies of electromagnetic radiation.
  - The Rutherford model is unable to explain this phenomena.
- Rutherford's electrons are undergoing a centripetal acceleration and so should radiate electromagnetic waves of the same frequency.
  - The radius should steadily decrease as this radiation is given off.
  - The electron should eventually spiral into the nucleus, but it doesn't.

Emission Spectra

- A gas at low pressure has a voltage applied to it.
- The gas emits light which is characteristic of the gas.
- When the emitted light is analyzed with a spectrometer, a series of discrete bright lines is observed.
  - Each line has a different wavelength and color.
  - This series of lines is called an *emission spectrum.*

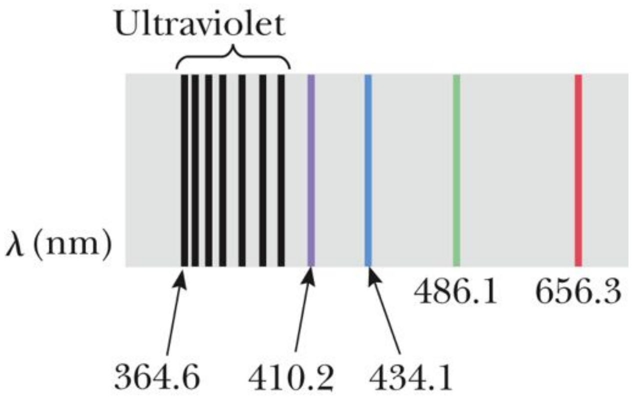
Examples of Emission Spectra



Emission Spectrum of Hydrogen – Equation

- The wavelengths of hydrogen's spectral lines can be found from $$\frac{1}{\lambda} = R_H\left(\frac{1}{2^2} - \frac{1}{n^2}\right)$$
  - R_H is the *Rydberg constant*
    - $R_H = 1.097\,373\,2 \times 10^7\ m^{-1}$
  - n is an integer, n = 1, 2, 3, …
  - The spectral lines correspond to different values of n.

Spectral Lines of Hydrogen



- The Balmer Series has lines whose wavelengths are given by the preceding equation.
- Examples of spectral lines
  - $n = 3$, $\lambda = 656.3$ nm
  - $n = 4$, $\lambda = 486.1$ nm

General Rydberg Equation

- The Rydberg equation can apply to any series.

$$\frac{1}{\lambda} = R_H \left(\frac{1}{m^2} - \frac{1}{n^2} \right)$$

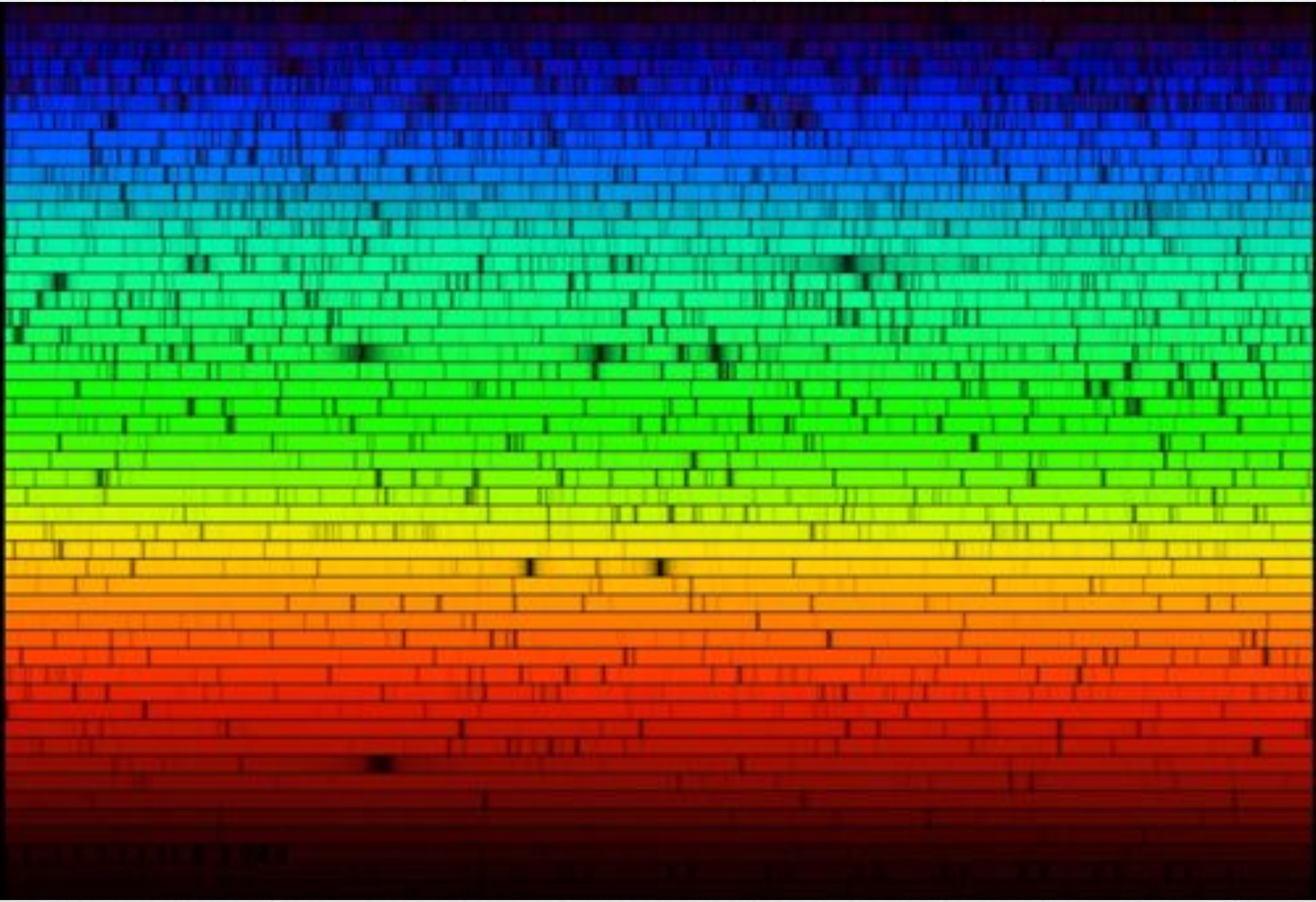
  - m and n are positive integers.
  - n > m.

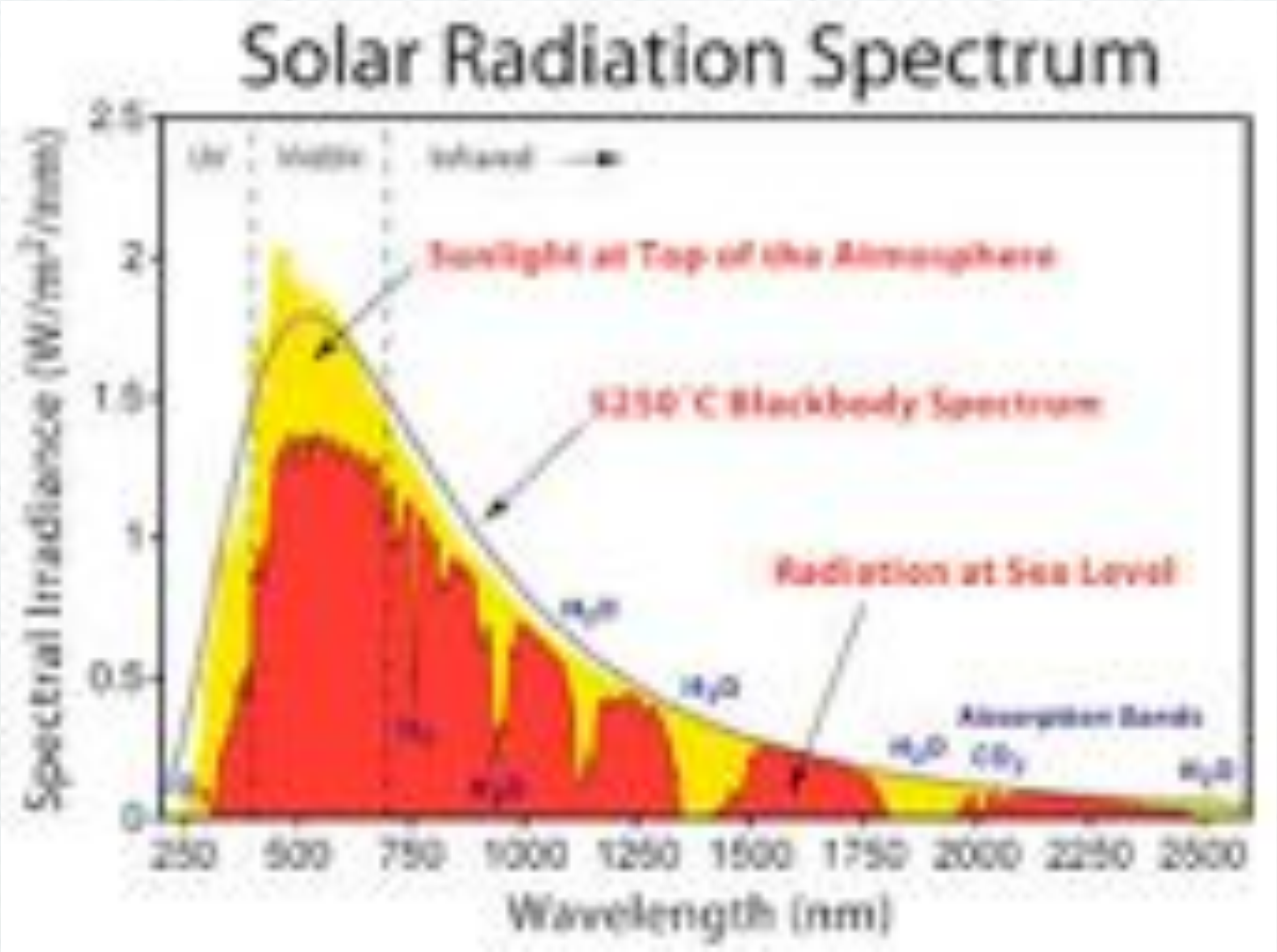
Absorption Spectra

- An element can also absorb light at specific wavelengths.
- An absorption spectrum can be obtained by passing a continuous radiation spectrum through a vapor of the element being analyzed.
- The absorption spectrum consists of a series of dark lines superimposed on the otherwise continuous spectrum.
  - The dark lines of the absorption spectrum coincide with the bright lines of the emission spectrum.

Application of Absorption Spectrum

- The continuous spectrum emitted by the Sun passes through the cooler gases of the Sun's atmosphere.
  - The various absorption lines can be used to identify elements in the solar atmosphere.
  - Led to the discovery of helium




Solar Radiation Spectrum
Sunlight at Top of the Atmosphere
5250°C Blackbody Spectrum
Radiation at Sea Level
Absorption Bands
Spectral Irradiance (W/m²/nm)
Wavelength (nm)
2.5
2
1.5
1
0.5
0
250
500
750
1000
1250
1500
1750
2000
2250
2500

Hw 2 & 5

Niels Bohr

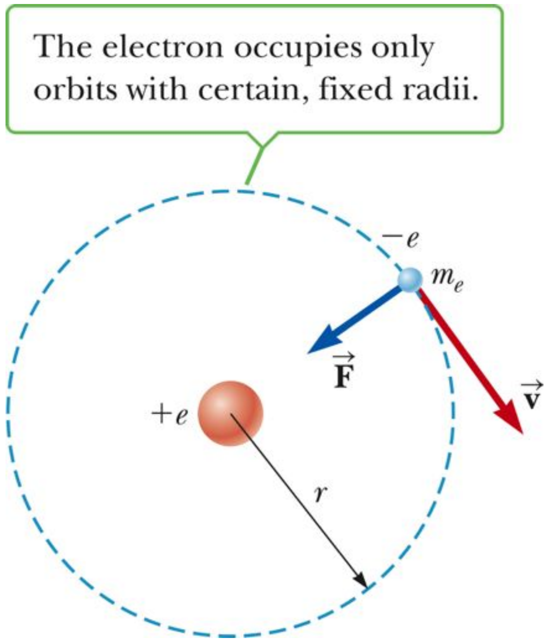
- 1885 – 1962
- Participated in the early development of quantum mechanics
- Headed Institute in Copenhagen
- 1922 Nobel Prize for structure of atoms and radiation from atoms

The Bohr Theory of Hydrogen

- In 1913 Bohr provided an explanation of atomic spectra that includes some features of the currently accepted theory.
- His model includes both classical and non-classical ideas.
- His model included an attempt to explain why the atom was stable.

Bohr's Assumptions for Hydrogen

- The electron moves in circular orbits around the proton under the influence of the Coulomb force of attraction.
  - The Coulomb force produces the centripetal acceleration.



Bohr's Assumptions, Cont.

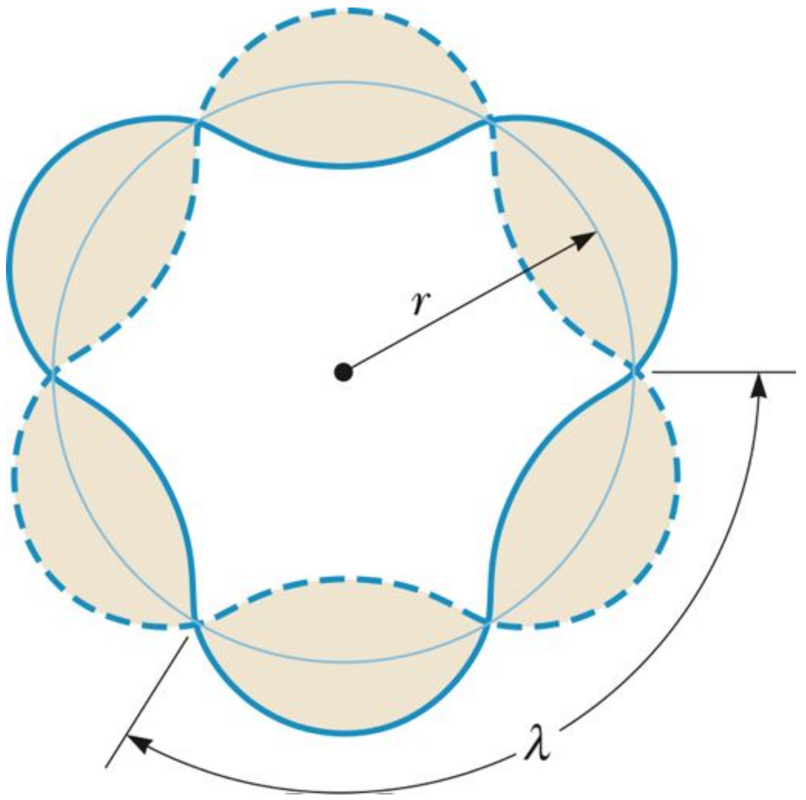
- Only certain electron orbits are stable and allowed.
  - These are the orbits in which the atom does not emit energy in the form of electromagnetic radiation.
  - Therefore, the energy of the atom remains constant.
- Radiation is emitted by the atom when the electron "jumps" from a more energetic initial state to a less energetic state.
  - The "jump" cannot be treated classically.

Bohr's Assumptions, Final

- The electron's "jump," continued
  - The frequency emitted in the "jump" is related to the change in the atom's energy.
  - The frequency is given by $E_i - E_f = h f$
  - It is *independent of the frequency of the electron's orbital motion.*
- The circumference of the allowed electron orbits is determined by a condition imposed on the electron's orbital angular momentum.

Electron' s Orbit

- The circumference of the electron' s orbit must contain an integral number of de Broglie wavelengths.
- $2 \pi r = n \lambda$
  - N = 1, 2, 3, …



Mathematics of Bohr's Assumptions and Results

- Electron's orbital angular momentum
  - $m_e v r = n \hbar$ where n = 1, 2, 3, …
- The total energy of the atom
  - $$E = KE + PE = \frac{1}{2} m_e v^2 - k_e \frac{e^2}{r}$$
- The energy of the atom can also be expressed as
  - $$E = -\frac{k_e e^2}{2r}$$

Bohr Radius

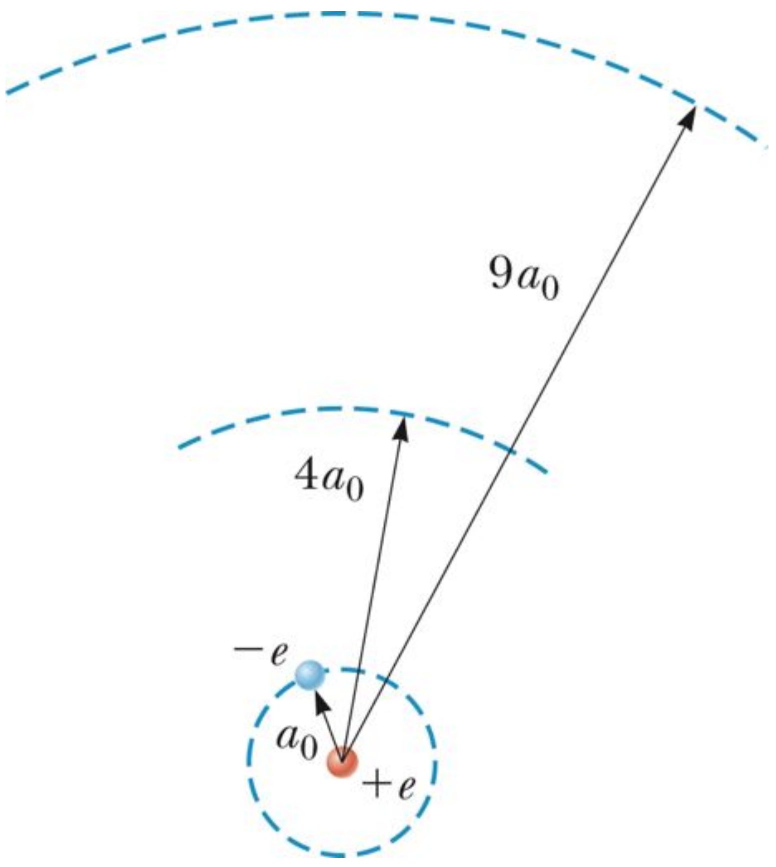
- The radii of the Bohr orbits are quantized.

$$r_n = \frac{n^2 \hbar^2}{m_e k_e e^2} \quad n = 1, 2, 3, \ldots$$

  - This is based on the assumption that the *electron can only exist in certain allowed orbits determined by the integer n.*
    - When n = 1, the orbit has the smallest radius, called the *Bohr radius*, a_o
    - a_o = 0.052 9 nm

Radii and Energy of Orbits

- A general expression for the radius of any orbit in a hydrogen atom is
  - $r_n = n^2 a_o$
- The energy of any orbit is
  - $E_n = -13.6\ eV/n^2$



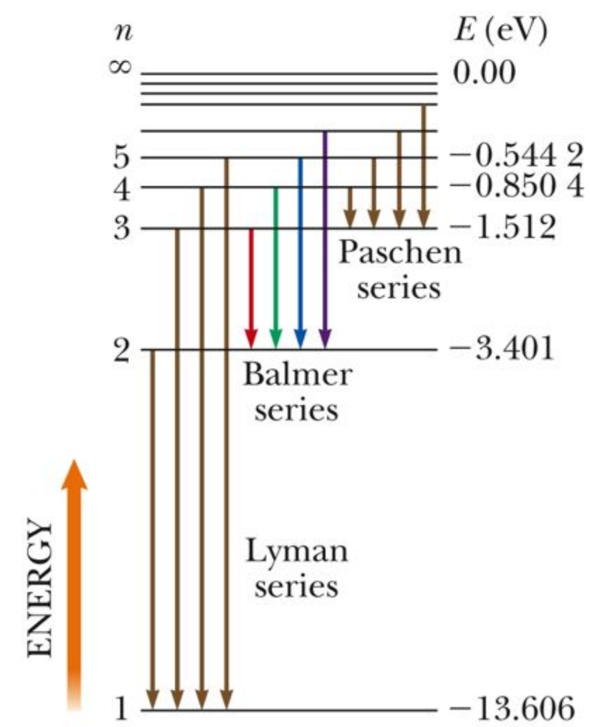
Specific Energy Levels

- The lowest energy state is called the *ground state.*
  - This corresponds to n = 1
  - Energy is –13.6 eV
- The next energy level has an energy of –3.40 eV.
  - The energies can be compiled in an *energy level diagram.*

Specific Energy Levels, Cont.

- The *ionization energy* is the energy needed to completely remove the electron from the atom.
  - The ionization energy for hydrogen is 13.6 eV
- The uppermost level corresponds to E = 0 and n → ∞

Energy Level Diagram



Generalized Equation

- The value of R_H from Bohr's analysis is in excellent agreement with the experimental value.
- A more generalized equation can be used to find the wavelengths of any spectral lines.

$$\frac{1}{\lambda} = R_H \left(\frac{1}{n_f^2} - \frac{1}{n_i^2} \right)$$

  - For the Balmer series, $n_f = 2$
  - For the Lyman series, $n_f = 1$
- Whenever a transition occurs between a state, n_i to another state, n_f (where $n_i > n_f$), a photon is emitted.
  - The photon has a frequency $f = (E_i - E_f)/h$ and wavelength λ

Hw 8, 13, & 15

Bohr's Correspondence Principle

- Bohr's *Correspondence Principle* states that quantum mechanics is in agreement with classical physics when the energy differences between quantized levels are very small.
  - Similar to having Newtonian Mechanics be a special case of relativistic mechanics when v << c

Successes of the Bohr Theory

- Explained several features of the hydrogen spectrum
- Can be extended to "hydrogen-like" atoms
  - Those with one electron
  - Ze^2 needs to be substituted for e^2 in equations.
    - Z is the atomic number of the element.

Quantum Mechanics and the Hydrogen Atom

- One of the first great achievements of quantum mechanics was the solution of the wave equation for the hydrogen atom.
- The energies of the allowed states are in exact agreement with the values obtained by Bohr when the allowed energy levels depend only on the principle quantum numbers.

Quantum Numbers

- n – principle quantum number
- Two other quantum numbers emerge from the solution of Schrödinger equation.
  - l – orbital quantum number
  - m_l – orbital magnetic quantum number

Quantum Number Summary

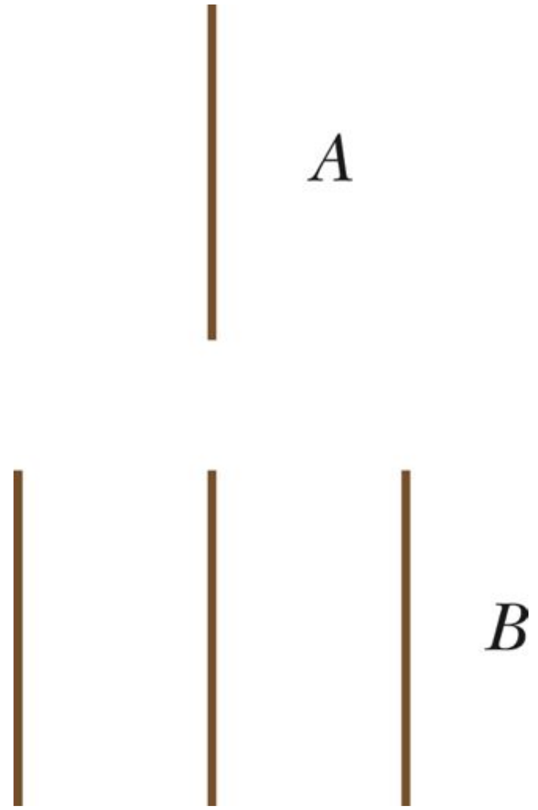
- The values of n can range from 1 to ∞ in integer steps.
- The values of ℓ can range from 0 to n-1 in integer steps.
- The values of m_ℓ can range from -ℓ to ℓ in integer steps.
  - Also see Table 28.1

Shells and Subshells

- All states with the same principle quantum number, n, are said to form a shell.
  - Shells are identified as K, L, M, …
  - Correspond to n = 1, 2, 3, …
- The states with given values of m and l are said to form a subshell.
- See table 28.2 for a summary.

Zeeman Effect

- The Zeeman effect is the splitting of spectral lines in a strong magnetic field.
  - This indicates that the energy of an electron is slightly modified when the atom is immersed in a magnetic field.
  - This is seen in the quantum number m_ℓ

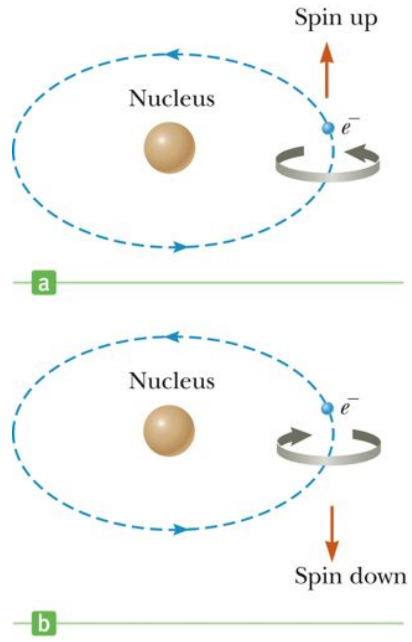


Spin Magnetic Quantum Number

- Some spectral lines were found to actually be two very closely spaced lines.
- This splitting is called **fine structure.**
- A fourth quantum number, spin magnetic quantum number, was introduced to explain fine structure.
  - This quantum number does not come from the solution of SchÖdinger's equation.

Spin Magnetic Quantum Number

- It is convenient to think of the electron as spinning on its axis.
  - The electron is *not* physically spinning.
- There are two directions for the spin
  - Spin up, $m_s = ½$
  - Spin down, $m_s = -½$
- There is a slight energy difference between the two spins and this accounts for the doublet in some lines.

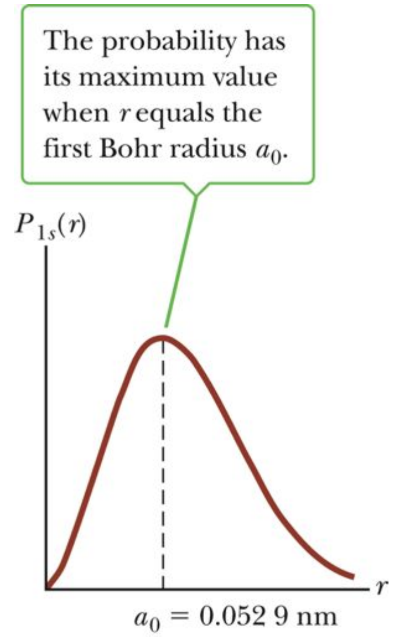


Spin Notes

- A classical description of electron spin is incorrect.
  - Since the electron cannot be located precisely in space, it cannot be considered to be a spinning solid object.
  - Electron spin is a purely quantum effect that gives the electron an angular momentum as if it were physically spinning.
- Paul Dirac developed a relativistic quantum theory in which spin naturally arises.

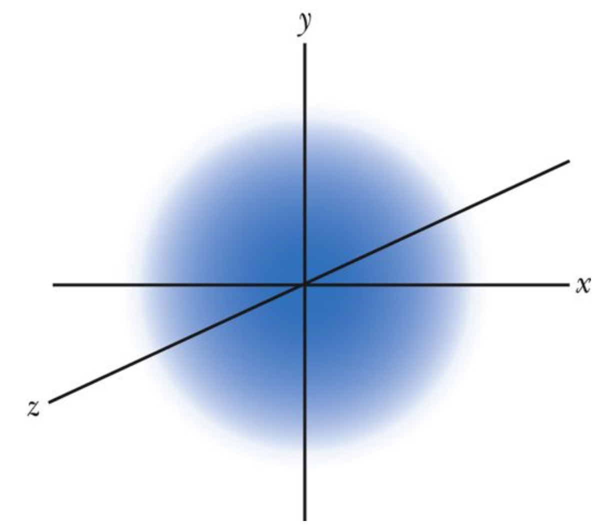
Electron Clouds

- The graph shows the solution to the wave equation for hydrogen in the ground state.
  - The curve peaks at the Bohr radius.
  - The electron is not confined to a particular orbital distance from the nucleus.
- The *probability* of finding the electron at the Bohr radius is a maximum.



Electron Clouds, Cont.

- The wave function for hydrogen in the ground state is symmetric.
  - The electron can be found in a spherical region surrounding the nucleus.
- The result is interpreted by viewing the electron as a cloud surrounding the nucleus.
  - The densest regions of the cloud represent the highest probability for finding the electron.



Wolfgang Pauli

- 1900 – 1958
- Contributions include
  - Major review of relativity
  - Exclusion Principle
  - Connect between electron spin and statistics
  - Theories of relativistic quantum electrodynamics
  - Neutrino hypothesis
  - Nuclear spin hypothesis

The Pauli Exclusion Principle

- No two electrons in an atom can ever have the same set of values of the quantum numbers n, ℓ, m_ℓ, and m_s
- This explains the electronic structure of complex atoms as a succession of filled energy levels with different quantum numbers.

Filling Shells

- As a general rule, the order that electrons fill an atom's subshell is:
  - Once one subshell is filled, the next electron goes into the vacant subshell that is lowest in energy.
  - Otherwise, the electron would radiate energy until it reached the subshell with the lowest energy.
  - A subshell is filled when it holds 2(2ℓ+1) electrons.
  - See table 28.3.

The Periodic Table

- The outermost electrons are primarily responsible for the chemical properties of the atom.
- Mendeleev arranged the elements according to their atomic masses and chemical similarities.
- The electronic configuration of the elements explained by quantum numbers and Pauli's Exclusion Principle explains the configuration.

Key Concepts

* Energy from light is quantized
* Light interacts with matter and vice versa
* Matter has a wavelike nature
* It is impossible to precisely know the position and momentum of a particle
* The orbit and energy of an electron is quantized

Key Equations

$$E_{photon} = h\nu$$

$$T_{peak} = \frac{.2898 \times 10^{-2} mK}{\lambda_{max}}$$

$$KE_{electron} = \frac{1}{2} m_{electron} V_{electron}^2$$

$$\Delta\lambda = \lambda - \lambda_0 = \frac{h}{m_e c}(1 - cos(\theta))$$

$$\lambda = \frac{h}{p}$$

$$\Delta x \Delta p_x \geq \frac{h}{4\pi}$$

Key Equations Cont

$$\frac{1}{\lambda} = R_H(\frac{1}{m^2} - \frac{1}{n^2})$$

$$\Delta E = hf$$

$$E = -\frac{k_e e^2}{2r}$$

$$r_n = \frac{n^2 \hbar^2}{m_e k_e e^2}$$

$$r_n = n^2 a_0 \rightarrow E_n = -13.6ev/n^2$$