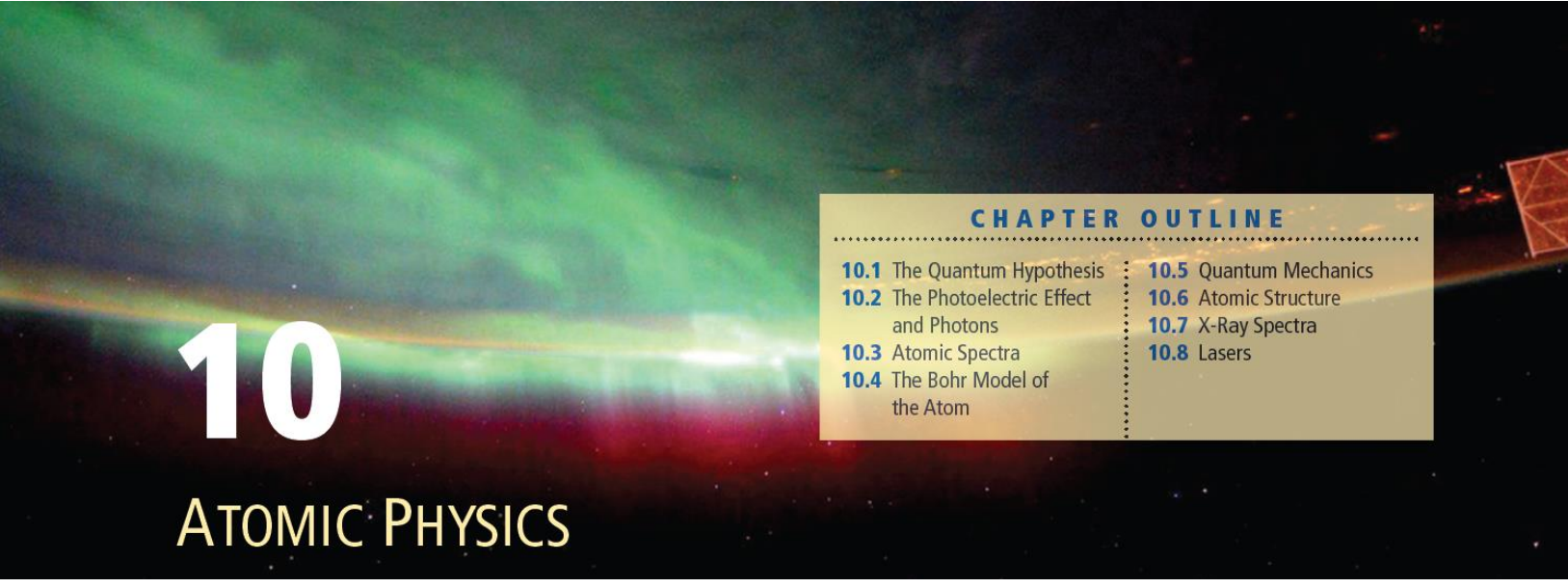

10

ATOMIC PHYSICS

CHAPTER OUTLINE

10.1 The Quantum Hypothesis
10.2 The Photoelectric Effect and Photons
10.3 Atomic Spectra
10.4 The Bohr Model of the Atom
10.5 Quantum Mechanics
10.6 Atomic Structure
10.7 X-Ray Spectra
10.8 Lasers

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CENGAGE

10.1 The Quantum Hypothesis

- By the end of the 1800's, physics had made significant progress.
  - Some physicists feared "that all of their questions might soon be answered."

- But some problems defied solution:
  1) blackbody radiation
  2) photoelectric effect
  3) atomic spectra

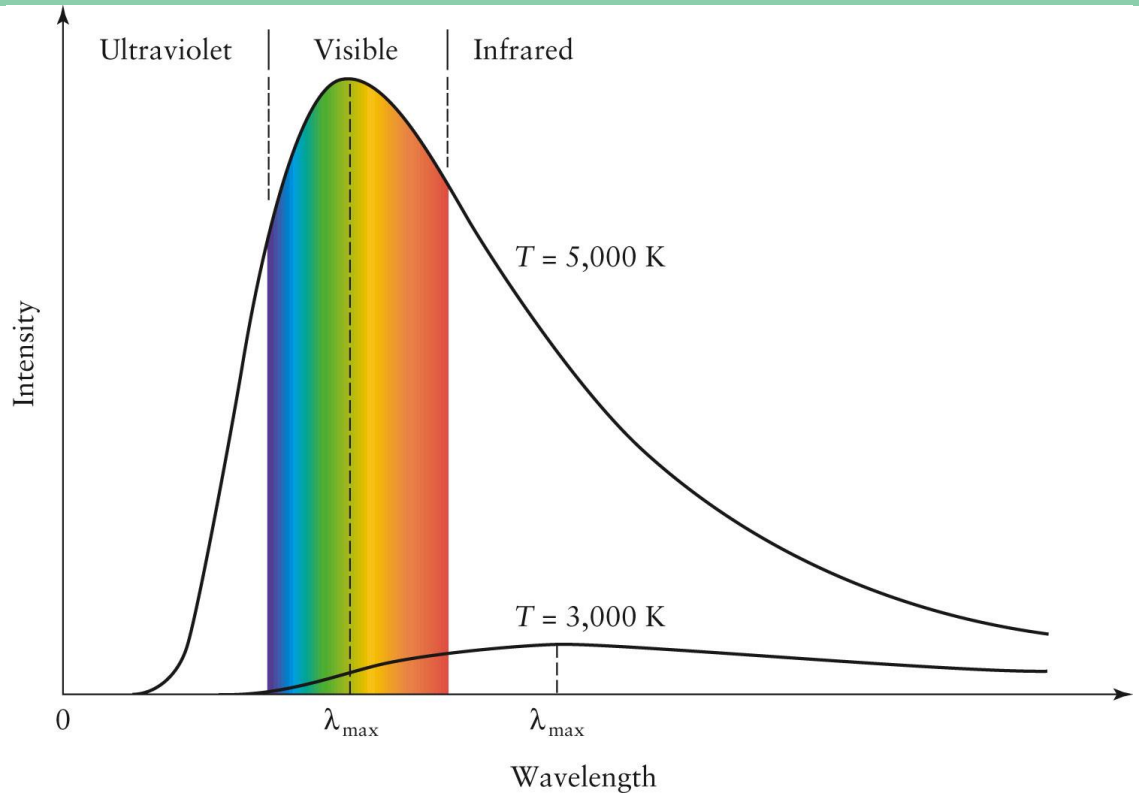
10.1 Blackbody Radiation

- Everything around you, as well as your own body, is constantly emitting electromagnetic (EM) radiation.
- At normal room temperature, mostly infrared (IR) radiation is emitted. Very hot bodies that are actually glowing, such as the Sun, emit visible light and ultraviolet (UV) radiation as well as IR.
- Bodies that are very cold, on the other hand, emit very weak, long-wavelength infrared and microwaves.

- This radiation had been studied carefully by scientists in the 19th century, and *its properties were well known but not understood*.

- For example, it had been determined that a perfectly "black" body, one that would absorb all EM radiation incident on it, would also be a perfect emitter of EM radiation. Such an ideal ***blackbody*** was said to emit ***blackbody radiation (BBR).***

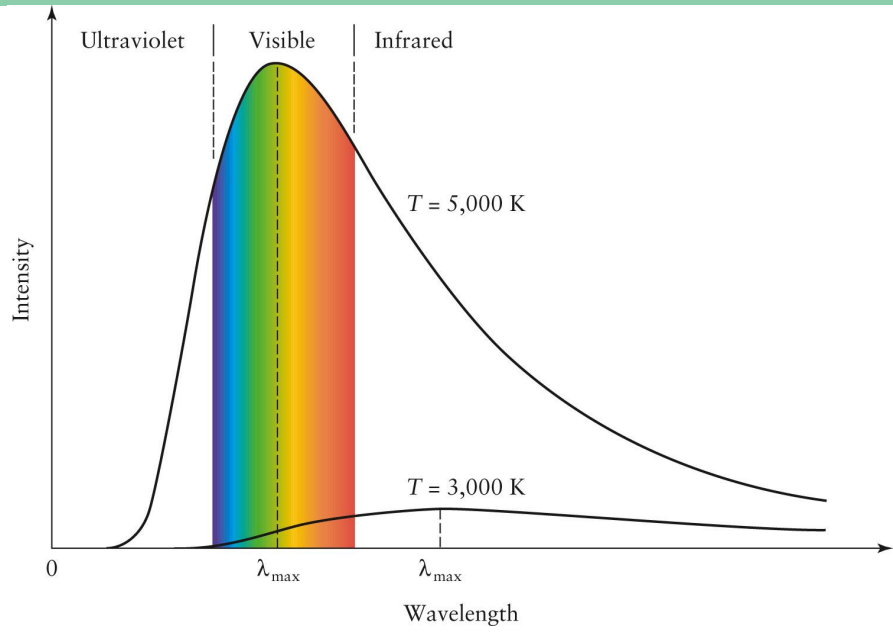
10.1 Blackbody Radiation



The BBR emitted by a given blackbody at a particular temperature can be plotted as the intensity of the radiation versus wavelength, is called a ***blackbody radiation curve***.

10.1 Blackbody Radiation

The graph illustrates two important ways that BBR changes *when the temperature of the body is increased*:



1) More energy is emitted per second at each wavelength of EM radiation.
2) The wavelength at which the most energy is emitted per second (in other words, the peak of the BBR curve) shifts to smaller values.

10.1 Blackbody Radiation

Why does a blackbody emit different amounts of EM radiation at different wavelengths in precisely this way?

- As a start, using the principles of electromagnetism, we can easily see why EM waves are emitted. Atoms and molecules are continually oscillating, and they contain charged particles—electrons and protons.

- You recall that that this kind of system will produce EM waves. But the exact mechanism involved, one that would account for the two observed features, *was a mystery* to scientists at the end of the 19th century.

10.1 Quantized Oscillators

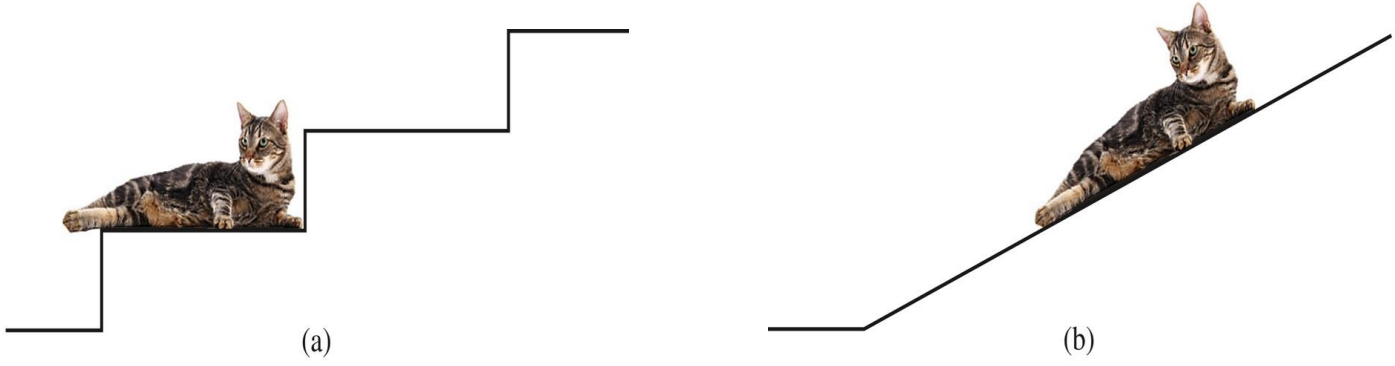
- The clue to solving this puzzle was discovered by German physicist Max Planck in the year 1900. Planck derived a mathematical equation that fit the shape of the BBR curves.

- Planck proposed that ***an oscillating atom in a blackbody can have only certain fixed values of energy***. It can have zero energy, or a particular energy E, or an amount equal to 2E, or 3E, or 4E, and so on.
- In other words, the energy of each atomic oscillator is ***quantized***.

- The energy E is called the fundamental quantum of energy for the oscillator. The allowed values of energy for the atom are integral multiples of this quantum.

$$\text{allowed energy} = nE, \text{ where } n = 0, 1, 2, 3, \ldots$$

10.1 Blackbody Radiation



- We can illustrate the difference between quantized and continuous energy values by comparing a stairway to a ramp.
- A cat lying on a stairway has quantized potential energy: it can be on only one of the steps, and each step corresponds to a particular PE.
- A cat lying on a ramp does not have quantized potential energy. It can be anywhere on the ramp. Its height above the ground can have any value within a certain range, and therefore its PE is one of a continuous range of values.

10.1 Blackbody Radiation

- Planck showed that light emitted in bursts by a blackbody, which resulted in the correct BBR curves.

- He determined that the oscillator's energy E, is proportional to frequency f.

$$E = \left(6.63 \times 10^{-34}\right) f = hf$$

$$\text{allowed energy} = nhf, \text{ with } n = 0, 1, 2, 3, \ldots$$

- The constant h is called Planck's constant and has SI units of joule-seconds.

10.2 The Photoelectric Effect

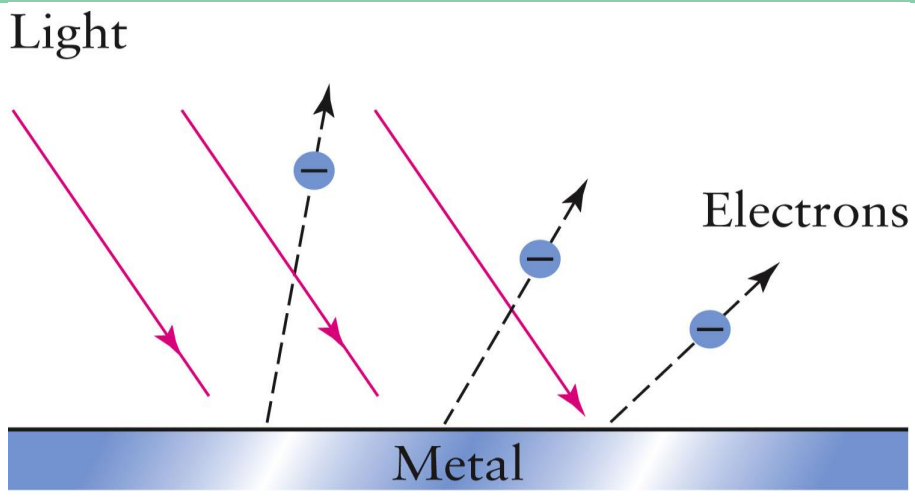
- The second phenomenon that puzzled physicists at the beginning of the 20th century was ***the photoelectric effect***.

- The photoelectric effect was accidentally discovered during experiments with EM waves.

- It was noticed that when metals were illuminated with UV light, intense sparks generated between two conductors, producing strong EM wave.

- The UV was somehow increasing the current in the spark.

10.2 Photoelectric Physics



- The ***photoelectric effect*** is exhibited by metals exposed to x-rays, UV light, or high-frequency visible light.
- The EM waves give energy to electrons in the metal, and the electrons are ejected from the surface.
- The biggest puzzle was the relationship between the speed or energy of the electrons and the incident light.
- Even more of a surprise was the finding that only the frequency (color) of the light affected the electron energies.

10.2 Photoelectric Physics

- The explanation of the photoelectric effect was supplied by Einstein.
- Planck had suggested that light is emitted in discrete "bursts" or "bundles." Einstein took this idea one step further and proposed that the light itself remains in bundles and is absorbed in this form.
- The electrons in a metal can take in light energy only by absorbing one of these discrete quanta of radiation.
- The amount of energy in each quantum of light depends on the frequency.
- This is the same equation for the energy of Planck's quantized atomic oscillators. Einstein suggested that light and other EM waves are quantized, just like the energy of oscillating atoms.

$$E = hf \quad \left(\text{energy of a quantum of EM radiation}\right)$$

10.2 Photoelectric Physics

- This idea, that the energy of an EM wave is quantized, allows us to picture the radiation as being made up of individual particles now called ***photons***.

- Each photon carries a quantum of energy, $E = hf$.

- The total energy in an EM wave is just the sum of the energies of all of the photons associated with it.

- Under certain circumstances, such as those involving refraction or interference of light, the wave nature of light is manifested.

- In other circumstances, including those involving the emission or absorption of light, the particle aspects of light are demonstrated.

10.2 Photoelectric Physics

With Einstein's proposal, the observed aspects of the *photoelectric effect could be explained*:

- *Higher-frequency light* ejects electrons with *more energy* because each photon has more energy to give.

- *Brighter light* simply means that *more photons strike* the metal each second. This results in *more electrons being ejected* each second, but it does not increase the energy of each electron.

10.2 Photoelectric Physics

How much energy does a typical photon have? Not very much.

- On this tiny scale, it is convenient to use a much smaller unit of energy called the electronvolt (eV).

- One electronvolt is the potential energy of each electron in a 1-volt battery.

10.2 Photoelectric Physics

- Because voltage is energy per charge, the charge of an electron multiplied by voltage equals energy.

$$\begin{aligned} 1 \text{ electronvolt} &= 1.6\times10^{-19} \text{ coulomb}\times 1 \text{ volt} \\ &= 1.6\times10^{-19} \text{ coulomb}\times 1 \text{ joule/coulomb} \\ 1 \text{ eV} &= 1.6\times10^{-19} \text{ J} \end{aligned}$$

- In terms of the electronvolt, the value of Planck's constant is:

$$h = 6.63\times10^{-34}\ \text{J}\cdot\text{s} = 4.14\times10^{-15}\ \text{eV/Hz}$$

Example 10.1

Compare the energies associated with a quantum of each of the following types of EM radiation.

$$\text{red light: } f = 4.3\times10^{14}\ \text{Hz}$$

$$\text{blue light: } f = 6.3\times10^{14}\ \text{Hz}$$

$$\text{x-ray: } f = 5\times10^{15}\ \text{Hz}$$

$$E = hf = 4.136\times10^{-15}\ \text{eV/Hz}\times4.3\times10^{14}\ \text{Hz}$$

$$E = 1.78\ \text{eV}\quad(\text{red light})$$

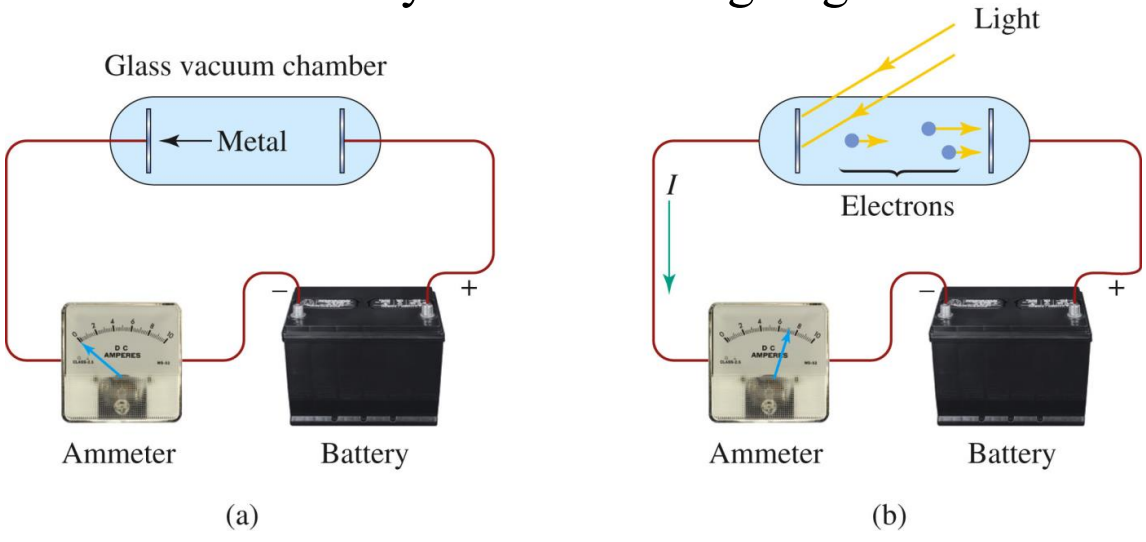
$$E = 2.61\ \text{eV}\quad(\text{blue light})$$

$$E = 20{,}700\ \text{eV}\quad(\text{x-ray})$$

Notice how much greater the energy of x-ray is compared to that of visible light. Because of that high doses of x-rays can be harmful to the human body.

10.2 Applications of the Photoelectric Effect

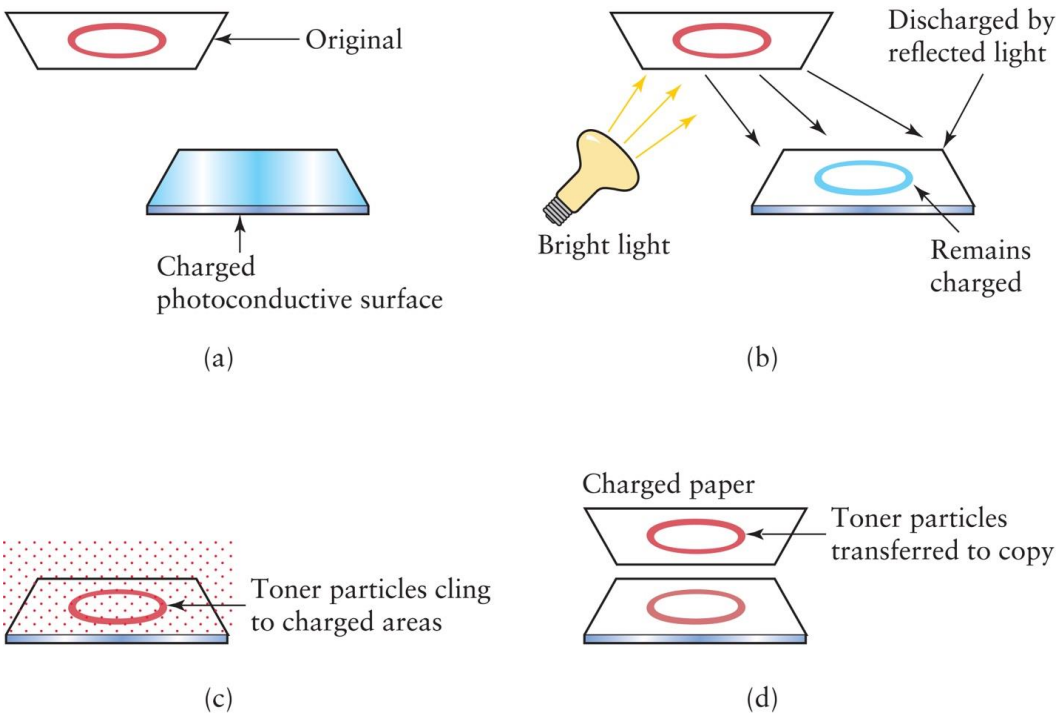
This phenomenon is the key to "interfacing" light with electricity.



- The figure shows a schematic of a device that can detect light.
- When no light strikes the metal, no current flows in the circuit because there is nothing to carry the charge through the tube.
- When light strikes the metal, electrons are ejected and attracted to the positive terminal of the tube. The result is a current flowing in the circuit.

10.2 Applications of the Photoelectric Effect

Photocopying machines use the interplay between electricity and light. The key part of the operation is a special photoconductive surface. It is normally an insulating material, but it becomes a conductor when exposed to light. Electrons bound to atoms are freed when they absorb photons in the incident light.



10.3 Atomic Spectra

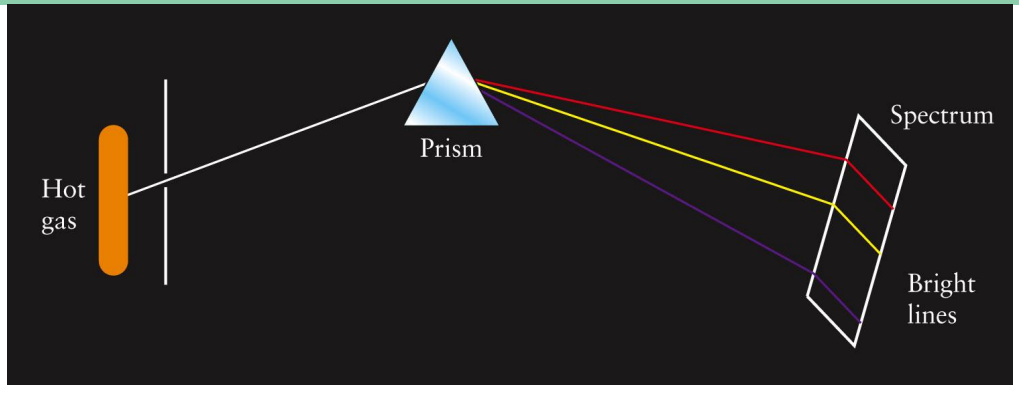
The third problem that defied explanation involved *the spectra produced by the various chemical elements*.

- Suppose we use a prism to examine the light produced by the heated filament of an ordinary incandescent bulb. The light from the bulb will be dispersed on passing through the prism.

- The spectrum will appear as a continuous band of the colors of the rainbow, one color smoothly blending into the next.

- Such a spectrum is called a *continuous spectrum.*

10.3 Atomic Spectra

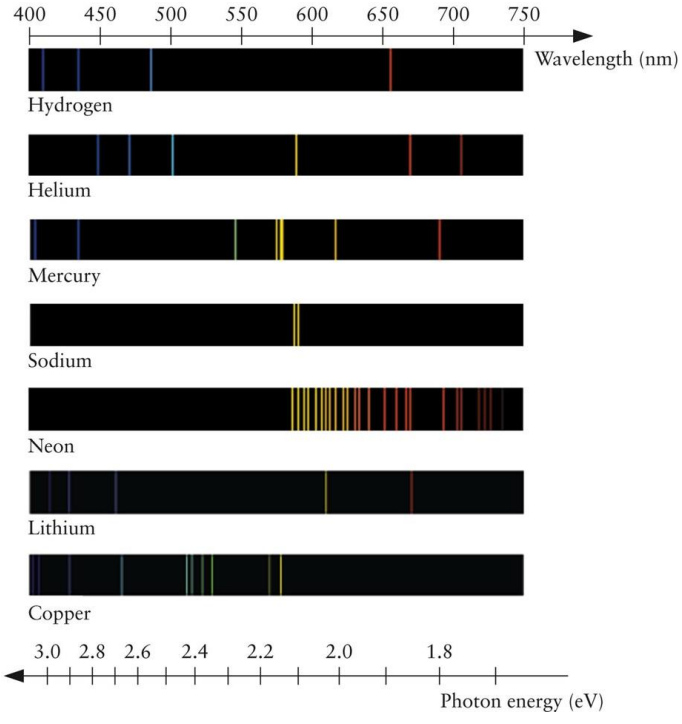


Imagine now a sample of some gas confined in a glass tube and heated to emit light.

- If we examine the light from the luminous gas with a prism, we do not see a continuous spectrum. Instead, we see an ***emission-line spectrum*** consisting of a few isolated, discrete lines of color.
- In this case, the source is not emitting radiation at all wavelengths (colors) but only at certain selected wavelengths.
- If a different type of gas is investigated, we find that the resultant line spectrum is different.

10.3 Atomic Spectra

Each type of gas has its own unique set of spectral lines.



The problem that remained was to understand why luminous gases produce line spectra and not continuous spectra, and how it is that each element has its own unique spectral "fingerprint" by which it can be identified.

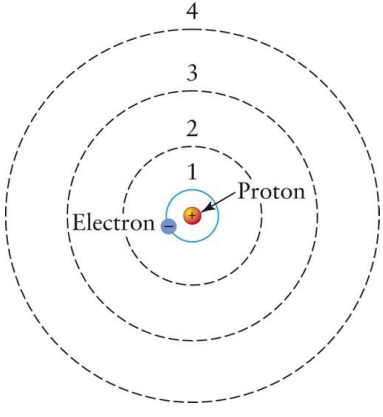
10.4 The Bohr Model of the Atom

- At the beginning of the 20th century, little was known about the atom. This made the origin of atomic spectra a mystery.
- In 1911, an important experiment performed by Rutherford revealed that the positive charge in an atom is concentrated in a tiny core—the nucleus.
- This result was quickly followed in 1913 by a model of the atom put forth by the physicist Niels Bohr.

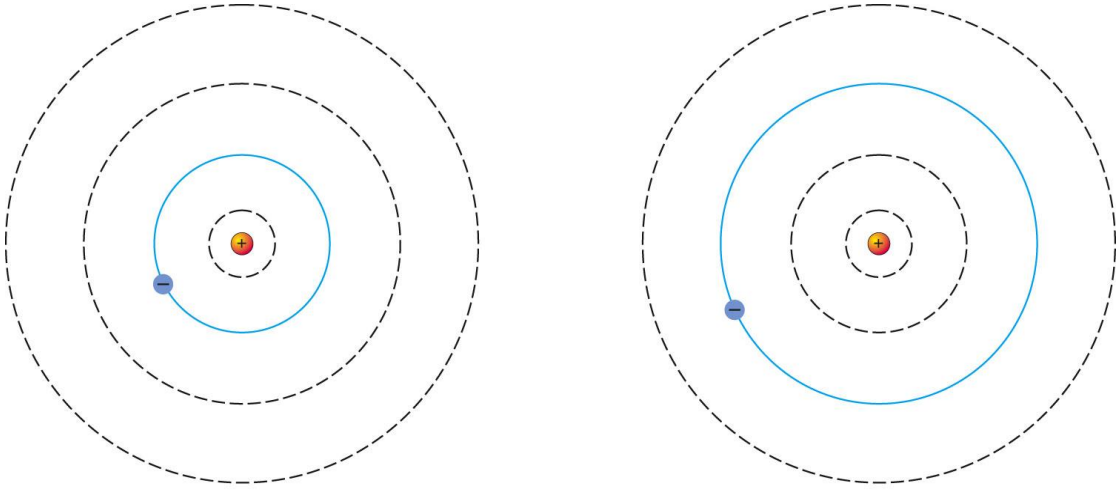
10.4 The Bohr Model of the Atom

Bohr's model of the atom, successfully explained the nature of atomic spectra. *The basic features of Bohr's model are as follows*:

1. The atom forms a miniature "solar system."
2. The electron orbits are quantized. Each allowed orbit has a particular energy associated with it, and the larger the orbit, the greater the energy.
3. Electrons do not radiate energy while in a stable orbit. Electrons may "jump" from one allowed orbit to another.



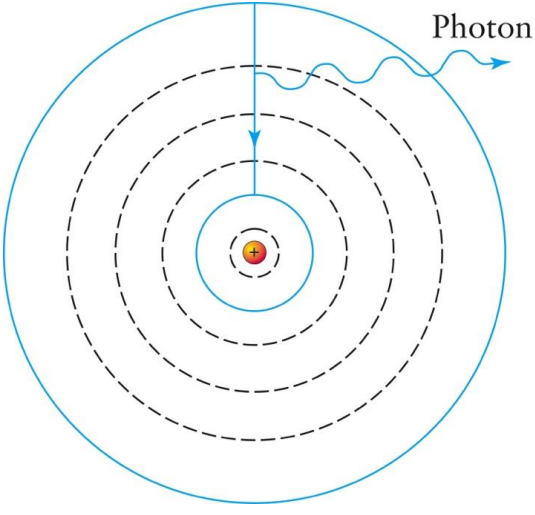
10.4 The Bohr Model of the Atom



- When in orbit 1, the electron has the lowest possible energy. The electron has more energy in each successively larger orbit.
- To go to a larger orbit, the electron must ***gain*** energy.
- The maximum energy the electron can have and still remain bound to the proton is called the ***ionization energy***.
- If the electron acquires more than this energy, it breaks free from the nucleus, and the atom is ionized.

10.4 Atomic Spectra Explained

How does the Bohr model account for the characteristic spectra emitted by luminous hydrogen gas?



- When an electron in a larger (higher-energy) orbit "jumps" to a smaller orbit, it ***loses*** energy.

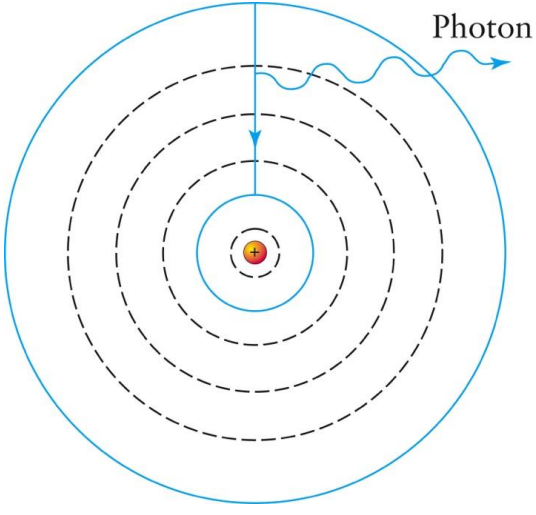
- One of the ways it can lose this energy is by *emitting light (photons)*. This process is the origin of atomic spectra.

10.4 Atomic Spectra Explained

- Imagine that the electron in the hydrogen atom is in the sixth allowed orbit, with its energy represented by E_6.

- Suppose the electron makes a transition (a jump) to an inner orbit—say, the second one—where its energy is E_2.

- To do so, the electron must lose an amount of energy, ΔE, equal to the difference in the two energy levels.



$$\Delta E = E_6 - E_2$$

$$\text{photon energy} = \Delta E = hf$$

10.4 Atomic Spectra Explained

In what is called a ***radiative transition***, the electron loses this energy through the emission of a photon that has an energy just equal to ΔE.

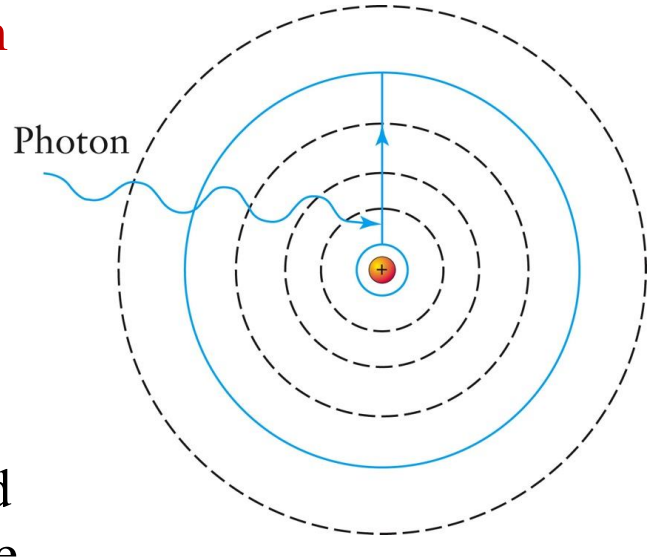
$$\text{photon energy} = \Delta E = hf$$

- The frequency of the emitted light is directly proportional to the difference in the energy of the orbits between which the electron jumped.
- The larger the energy difference, the higher the frequency of the light given off in the process.

$$\Delta E = E_3 - E_2 \text{ is smaller than } \Delta E = E_6 - E_2$$

10.4 Atomic Spectra Explained

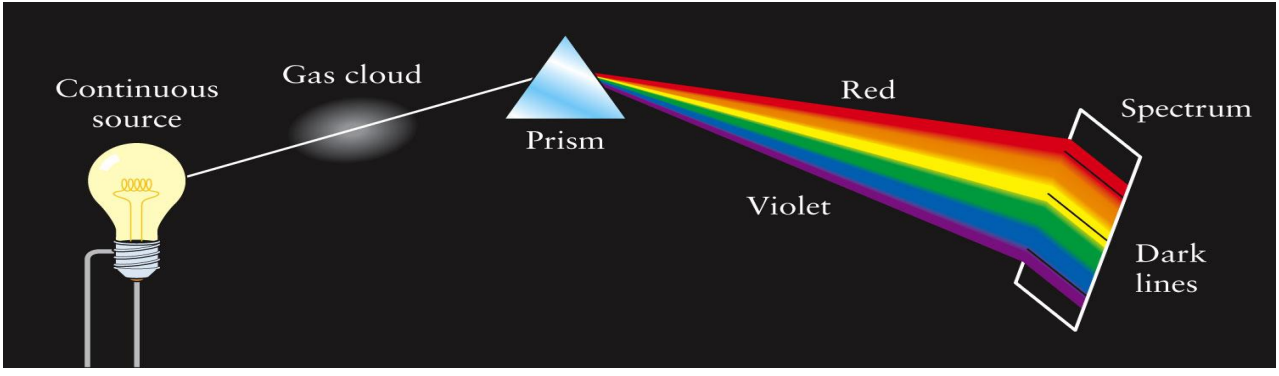
- The reverse of this process can also occur:
- An electron in an inner orbit can gain just the right amount of energy and jump to an outer orbit.
- For example, a hydrogen atom in the lowest energy state might gain the amount of energy needed for its electron to jump from orbit 1 to, say, orbit 5. To do this, the electron would have to acquire energy, as shown here.



$$\Delta E = E_5 - E_1$$

- One way an electron can jump to an outer orbit is by ***absorbing a photon*** with the proper energy. For example, an electron in orbit 1 can jump to orbit 5 if it absorbs a photon with energy ΔE.

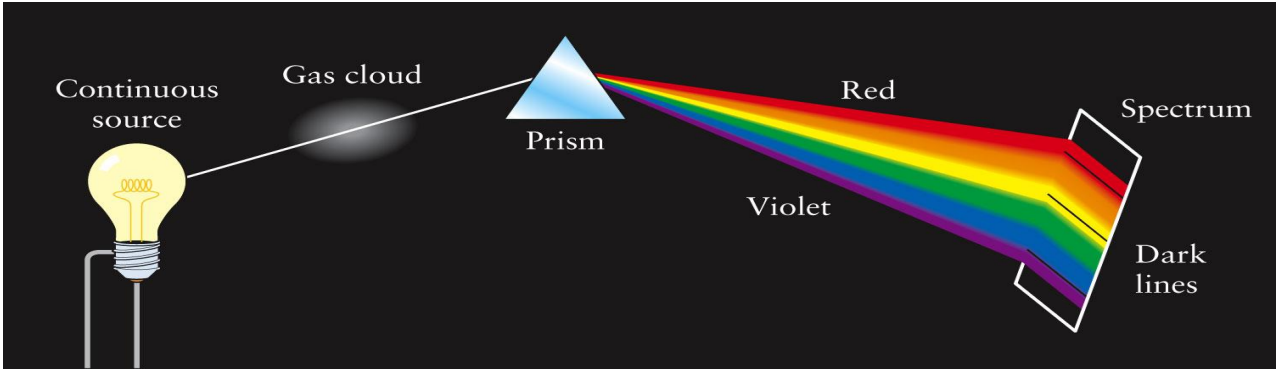
10.4 Atomic Spectra Explained



If white light is passed through hydrogen gas and then dispersed with a prism, many such transitions to outer orbits will occur, also dark bands appear at certain frequencies.

- Some of the photons in the incident radiation will have just the right energies to induce transitions from inner orbits to outer orbits.

10.4 Atomic Spectra Explained



- In the process, the number of photons with these particular energies will be reduced, and the intensity of the EM radiation at the corresponding frequencies will decrease.

- This reduction of intensity of light at only certain frequencies after it passes through a gas results in an ***absorption spectrum***.

10.4 Atomic Spectra

- Bohr's model of electron orbits assumed there are only certain allowed orbits, meaning the electron orbits were restricted to specific radii.

- These radii were determined by a condition: that the angular momentum of the electron in its orbit is quantized.

- Much like the energy of Planck's quantized atomic oscillators, the orbital angular momentum of Bohr's electrons could only be integer multiples of Planck's constant divided by 2π, shown here.

$$\text{allowed angular momentum} = n\left(\frac{h}{2\pi}\right), \text{ with } n = 1, 2, 3, \ldots$$

10.5 Quantum Mechanics

What goes on at the atomic level?

- De Broglie proposed that **electrons and other particles** possess ***wavelike properties***.
- Einstein had shown that light has both wavelike and particle-like properties, so why not electrons too?
- The wave associated with any moving particle has a specific wavelength (***the de Broglie wavelength***) that depends on the particle's momentum, mv.

$$\lambda = \frac{h}{mv} \quad \left(\text{de Broglie wavelength}\right)$$

- The higher the momentum of a particle, the shorter its wavelength. High-speed electrons have shorter wavelengths than low-speed electrons.

Example 10.2

What is the de Broglie wavelength of an electron with speed 2.19×10^6 m/s in the smallest orbit in the hydrogen atom?

$$mv = 9.11 \times 10^{-31}\,\text{kg} \times 2.19 \times 10^{6}\ \text{m/s}$$
$$= 1.995 \times 10^{-24}\ \text{kg-m/s}$$

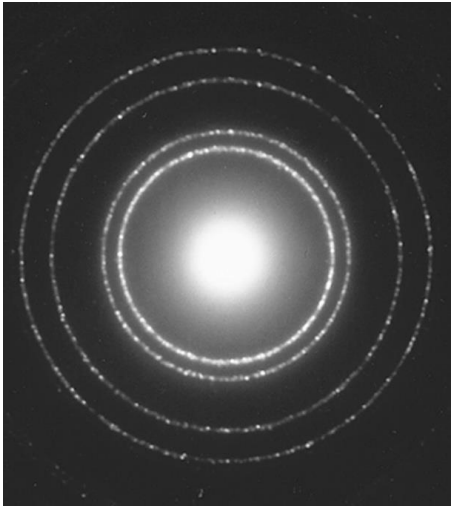
$$\lambda = \frac{h}{mv} = \frac{6.63 \times 10^{-34}\,\text{J-s}}{1.99 \times 10^{-24}\,\text{kg-m/s}}$$
$$= 3.32 \times 10^{-10}\,\text{m} = 0.332\ \text{nm}$$

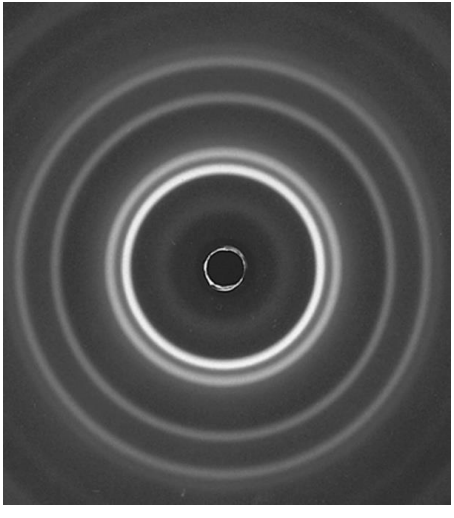
10.5 Quantum Mechanics

- Some experiments were interpreted as proof of the existence of de Broglie's matter waves.

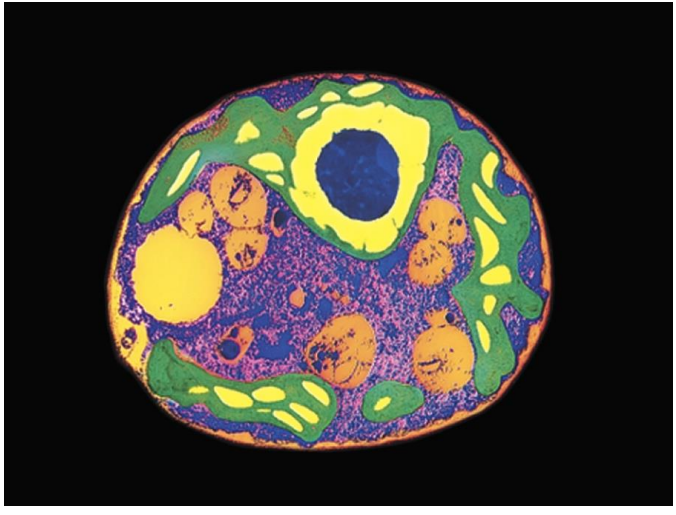
- These experiments showed that a beam of high-speed electrons underwent diffraction when sent into a metal target.

- The electrons were behaving just like waves. In fact, we get the same kind of scattering pattern using x-rays or electrons.





10.5 Quantum Mechanics



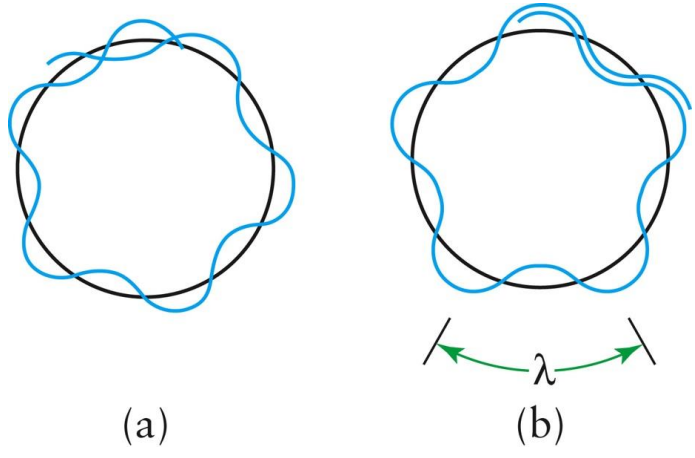
- The magnification can be much higher with an electron microscope because the de Broglie wavelength of the electrons is much shorter than the wavelengths of visible light.
- Diffraction and interference of waves passing around and between small objects such as cells strongly affect image clarity.
- Shorter-wavelength waves (electrons) diffract and interfere much less than do longer-wavelength waves (light).

10.5 The Bohr Atom Explained

- Another major triumph for de Broglie's wave hypothesis was its ability to explain Bohr's quantized orbits.

  - De Broglie reasoned that because an orbiting electron acts like a wave, its wavelength has to affect the circumference of its orbit.

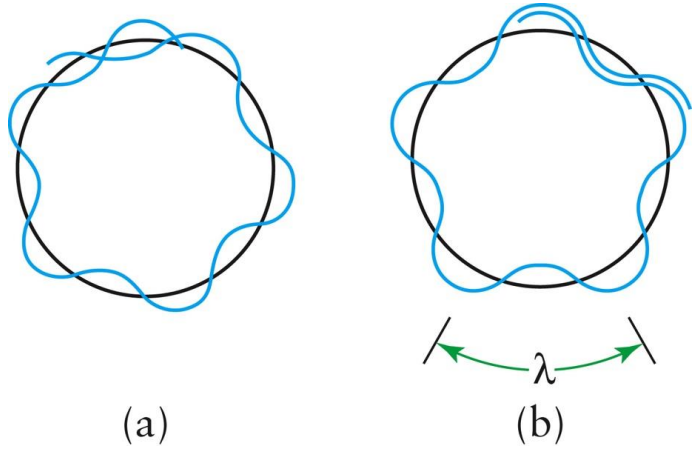
  - In particular, the electron's wave wraps around on itself, and in doing so it must interfere with itself.



10.5 The Bohr Atom Explained

- Only if the wave *interferes constructively* (peak matches peak) can the electron's orbit remain stable.

- This means that the circumference of its orbit must *equal* exactly one de Broglie wavelength, or two, or three, and so on.



- If r is the radius of a circular orbit, then the circumference is $2\pi r$. Thus, the radius is given by $n\lambda$ over 2π.

$$\text{circumference of orbit} = n\lambda, \text{ with } n = 1, 2, 3, \ldots$$

$$r = n\left(\frac{\lambda}{2\pi}\right), \text{ with } n = 1, 2, 3, \ldots$$

Example 10.3

Using the results of Example 10.2, find the radius of the smallest orbit in the hydrogen atom.

$$r = \frac{\lambda}{2\pi}$$

$$r = \frac{0.332 \text{ nm}}{2\pi} = \frac{0.332 \text{ nm}}{2 \times 3.14} = \frac{0.332 \text{ nm}}{6.28}$$

$$= 0.0529 \text{ nm}$$

10.5 The Bohr Atom Explained

- The allowed values for the circumference (or radius) of an electron's orbit leads to Bohr's allowed values for the electron's angular momentum.
  - Because the de Broglie wavelength is given by the equation here, the radius can have these values.

$$r = n\left(\frac{\lambda}{2\pi}\right),\ \lambda = \frac{h}{mv}$$

$$r = \frac{n}{2\pi}\left(\frac{h}{mv}\right),\ \text{with } n = 1,\ 2,\ 3,\ldots$$

$$mvr = n\left(\frac{h}{2\pi}\right),\ \text{with } n = 1,\ 2,\ 3,\ldots$$

The quantity *mvr* is the angular momentum of the electron.
De Broglie's condition on the size of the electron's orbit turns out to be identical to Bohr's condition on the angular momentum of the electron.

10.5 The Bohr Atom Explained

- Because electrons and other particles on the atomic scale have wavelike properties, we can no longer think of them as being like very tiny marbles or ball bearings.

- Rather, they are a bit spread out in space, more like tiny, fuzzy cotton balls.

- In the old particle model of the electron, it was possible to state exactly where an electron is and exactly what its momentum is at any instant in time.

- ***Heisenberg*** stated that the wave nature of particles makes this impossible. ***You cannot specify both the position and the momentum of an electron to arbitrarily high precision***.

10.5 The Bohr Atom Explained

- The more precisely you know the position of the electron, the less precisely you can determine its momentum and vice versa.

- If we let Δx represent *the uncertainty in the position of a particle* and Δmv represent the uncertainty in the momentum of the particle, then we get the ***uncertainty principle*** here: The product of the two uncertainties is greater than or equal to Planck's constant.

$$\Delta x \Delta mv \geq h \quad \left(\text{uncertainty principle}\right)$$

- No matter how good the experimental apparatus, the best precision we can achieve in the measurement of a particle's position and momentum is limited by this principle.
- On the atomic scale, particles cannot be localized in the same way that they can be on a large scale.

10.5 The Bohr Atom Explained

Schrödinger established *a mathematical model for the waves associated with particles*.

- In his model, a quantum system like a hydrogen atom can be described by a wave function.

- The wave function contains all the information needed to predict the characteristics and future evolution of the system once the nature of the interactions in which the system is involved are known.

- For example, knowledge of the wave function allows us to calculate the most probable position of the electron in a hydrogen atom with respect to the nucleus, as well as the electron's average energy.

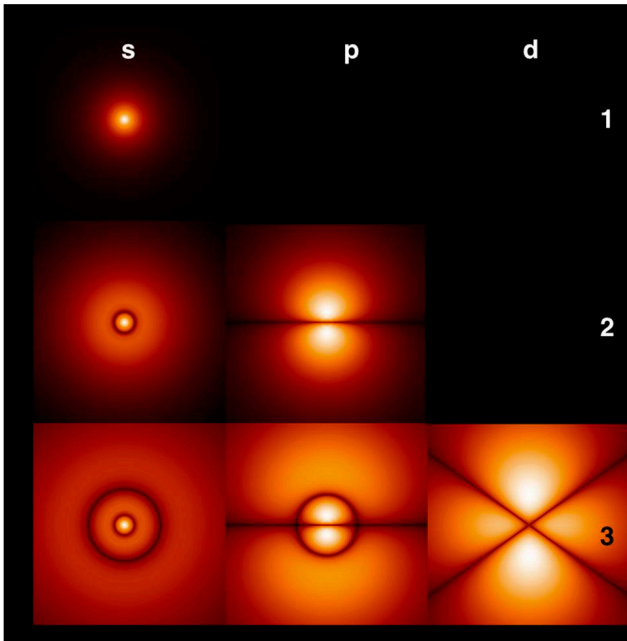
10.6 Atomic Structure

- We can no longer represent electrons as little particles with well-defined boundaries moving in precisely defined circular orbits.

- In reality, each electron is described by a wave that determines the probability of finding it at different locations.

- The atom is pictured as a tiny nucleus surrounded by an "electron cloud."

- The density of the cloud at each point in space indicates the likelihood of finding the electron there. The different allowed orbits of the electrons appear as clouds with different sizes and shapes.

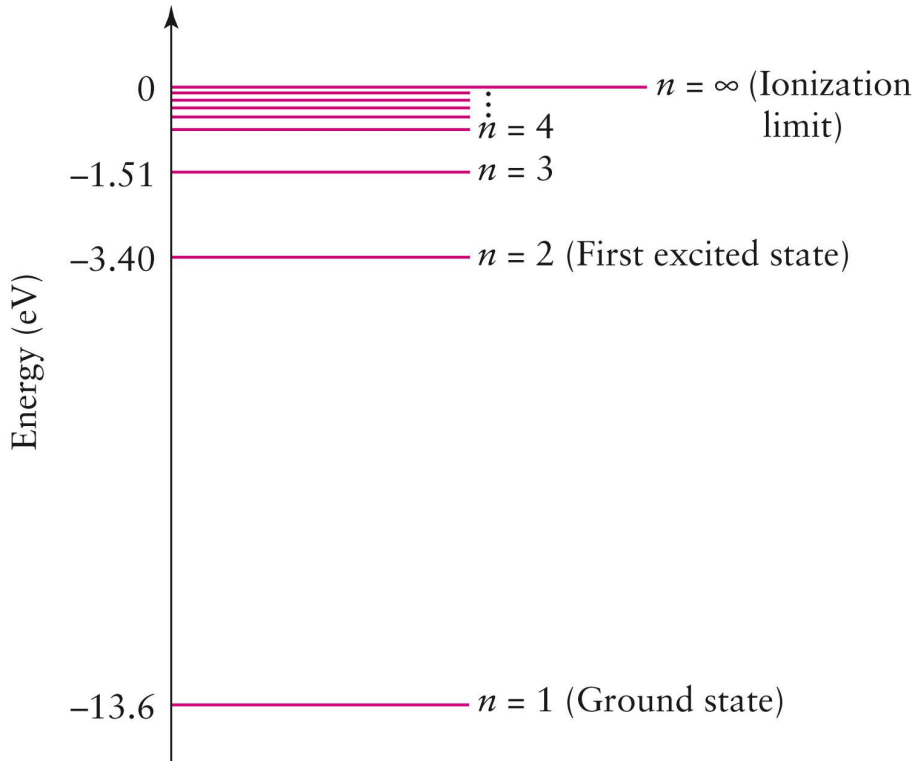


10.6 Atomic Energy Levels

- In the new model, we describe the electrons as being in certain allowed energy states or energy levels.

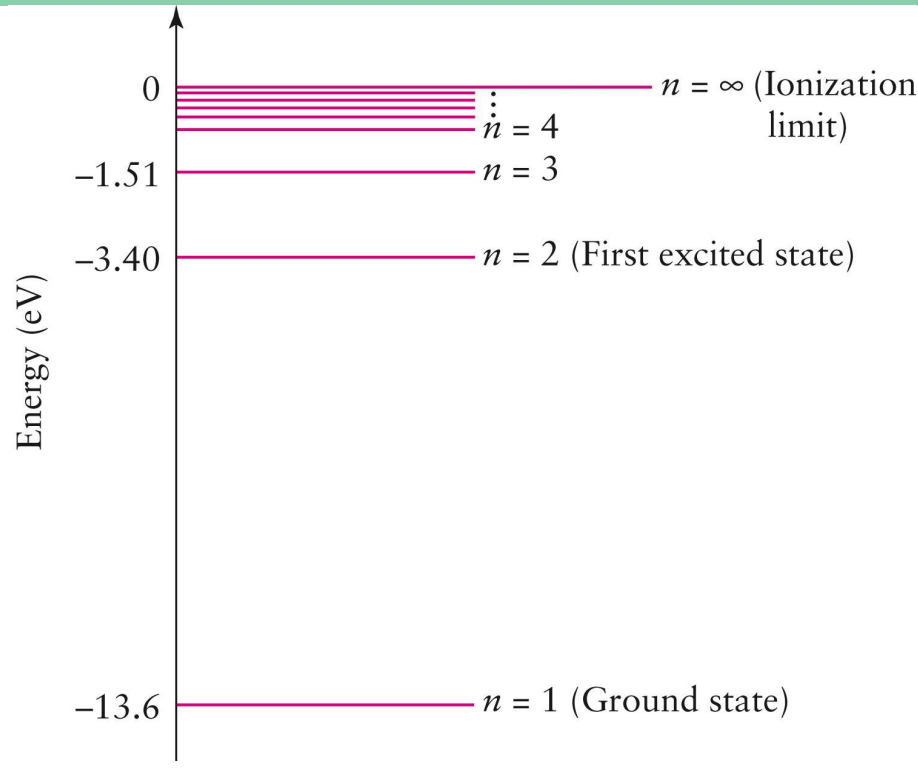
- For the hydrogen atom, the ***lowest energy level*** (corresponding to the innermost orbit in the Bohr theory) is called ***the ground state***.

- The ***higher energy states*** are referred to as ***excited states***.



10.6 Atomic Energy Levels

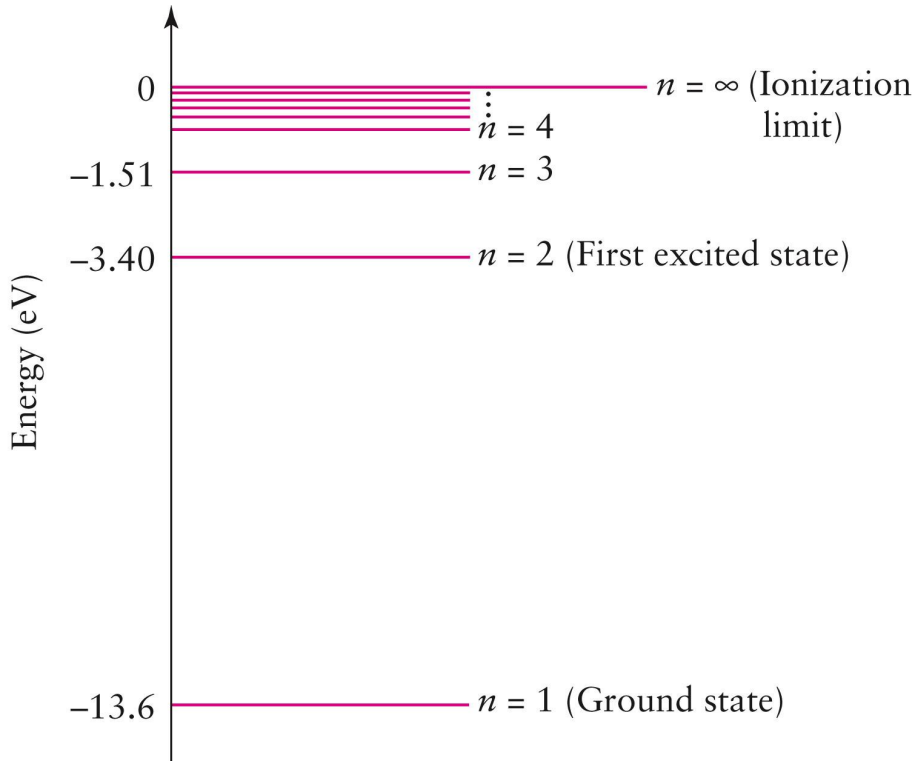
- We now represent the structure of the atom schematically using an ***energy-level diagram***.
- Each energy level is labeled with a quantum number, ***n***, beginning with the ground state having n = 1 and continuing on up.
- As the quantum number increases, so does the energy associated with the state.



- As *n gets larger*, the *difference in energy* between the adjacent states becomes *smaller*.

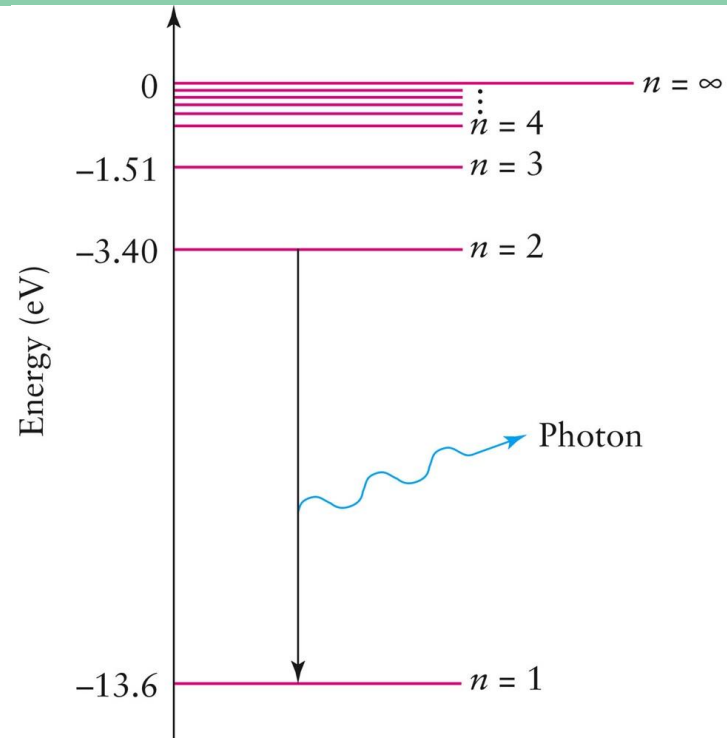
10.6 Atomic Energy Levels

- Note the existence of a *maximum allowed energy above which the electron is no longer bound to the nucleus*. The state is designated ***n* = ∞**, and its energy is the ***ionization energy***.
- The numbers on the left in the energy-level diagram are the electron energies for each state.
- The negative values indicate that the electron is bound to the nucleus.



10.6 Atomic Energy Levels

- Using the energy-level diagram, we can give a more complete picture of the emission spectrum of hydrogen.
- Suppose an atom in the n = 2 state undergoes a transition to the n = 1 state by emitting a photon.
- The transition is represented by an arrow drawn from the initial level to the final level.



- The photon that is emitted has an energy equal to the difference in energy between the two levels.

$$\text{photon energy} = \Delta E = E_2 - E_1 = -3.4\text{ eV} - (-13.6\text{ eV})$$
$$= 10.2\text{ eV}$$

Example 10.4

Find the frequency and wavelength of the photon emitted when a hydrogen atom goes from the $n = 3$ state to the $n = 2$ state.

We first find the energy of the photon and then use that to determine its frequency.

$$\text{photon energy} = hf = E_3 - E_2$$

$$= -1.51\text{ eV} - (-3.4\text{ eV}) = 1.89\text{ eV}$$

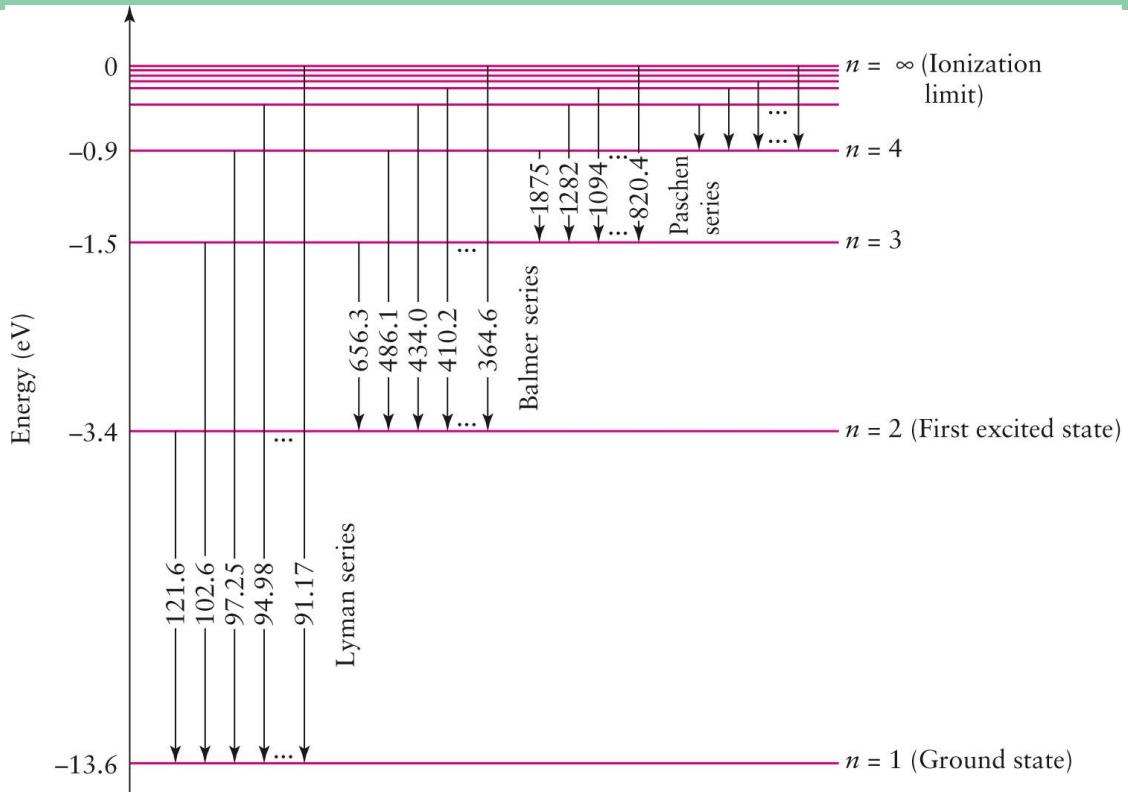
$$f = \frac{1.89\text{ eV}}{h} = \frac{1.89\text{ eV}}{4.136\times10^{-15}\text{eV/Hz}}$$

$$= 4.57\times10^{-14}\text{ Hz}$$

$$\lambda = \frac{c}{f} = \frac{3\times10^{8}\text{ m/s}}{4.57\times10^{14}\text{ Hz}}$$

$= 6.56 \text{ x } 10^{-7}$ m = 656 nm; This is a photon of visible light.

10.6 Atomic Structure



- Similarly, we can represent all possible downward transitions from higher energy levels to lower energy levels.
- The number with each arrow is the wavelength (in nanometers) of the photon that is emitted.

10.6 Atomic Structure

1. Transitions from higher energy levels **to the ground state *n* = 1** → emission of **UV** photon (***Lyman series***)

2. Transitions from higher energy levels **to *n* = 2** → emission of **visible** photons (***Balmer series***)

3. Transitions from higher energy levels **to *n* = 3** → emission of **infrared** photons (***Paschen series***)

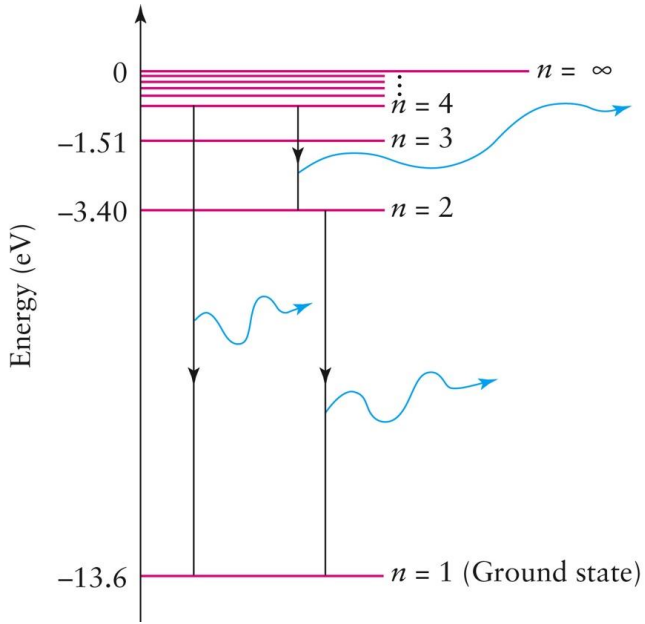
4. Downward transitions to other states with ***n* greater than 3**→ emission of other, **lower-energy photons**.

10.6 Atomic Structure

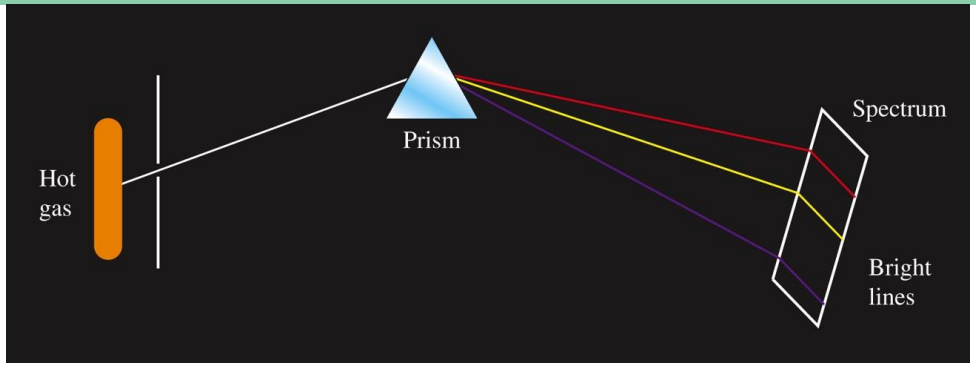
- An atom in the n = 3 state or higher can make transitions to intermediate energy levels instead of jumping directly to the ground state.

- For example, an atom in, say, the n = 4 state may jump to the n = 1 state, or it may go from n = 4 to n = 2 and then from n = 2 to n = 1.

- Atoms in the higher energy levels can undergo different "***cascades***" in returning to the ground state.

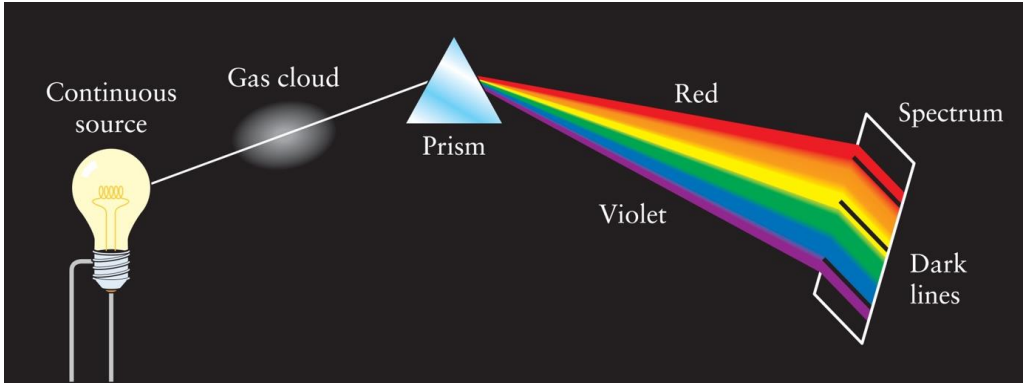


10.6 Atomic Structure



- *What actually happens when hydrogen gas in a tube is heated to a high temperature?*
- The billions and billions of atoms will be excited, some to each of the possible higher energy levels.
- The excited atoms then undergo spontaneous transitions to lower levels, with different atoms "stopping" at different levels.
- Thus, the hydrogen gas continuously emits photons corresponding to all of the possible energy-level transitions, although *some transitions are more favored than others*, depending in part on the gas temperature. This is hydrogen's ***emission spectrum***.

10.6 Atomic Structure



- Upward energy-level transitions occur when the atom gains energy by absorption of a photon with the proper energy.

- When white light—containing photons with many different energies—passes through hydrogen, those photons with just the right energy can be absorbed, this reduces the number of photons of the corresponding energy and frequency, appearing as dark bands.

- This leads to the observed ***absorption spectrum***.

10.6 Some Applications

Examples of the emission spectra of elements are the neon sign and the aurora borealis (northern lights).

10.6 The Pauli Exclusion Principle

The structure of atoms *with more than one electron* is governed by a principle formulated by Wolfgang ***Pauli***.

- He carefully analyzed the emission spectra of different elements, noting that some expected transitions did not occur when the lower energy level was already occupied by a certain number of electrons.
- From this and other evidence, *Pauli concluded that only a limited, fixed number of electrons could populate each energy level in an atom.*

PRINCIPLES

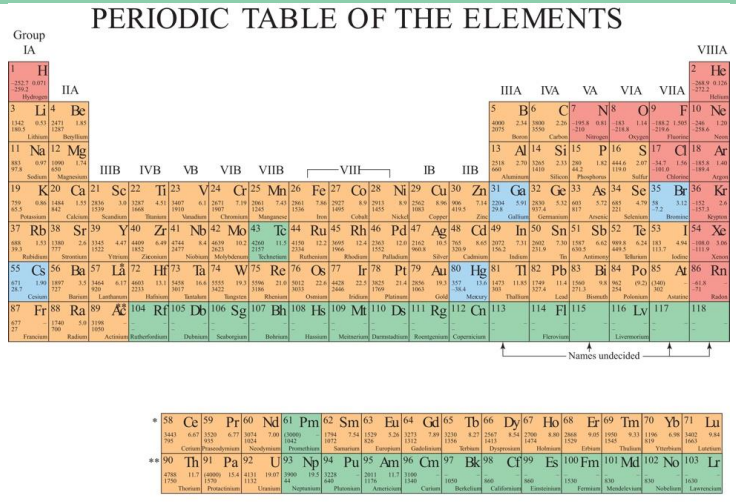
Pauli Exclusion Principle: Two electrons cannot occupy precisely the same quantum state at the same time.

10.6 The Pauli Exclusion Principle

- For each energy level, there exists a set number of quantum states available to the electrons.
- Once all of the quantum states corresponding to a given energy are filled, any remaining electrons in the atom must occupy other energy levels that have vacancies.
- *The general rule is, for level n, the occupation limit of electrons is* $2n^2$.

Element	Atomic Number	Number of Electrons in Level		
		$n = 1$	$n = 2$	$n = 3$
Hydrogen	1	1	0	0
Helium	2	2	0	0
Lithium	3	2	1	0
Carbon	6	2	4	0
Oxygen	8	2	6	0
Neon	10	2	8	0
Sodium	11	2	8	1

10.6 The Pauli Exclusion Principle



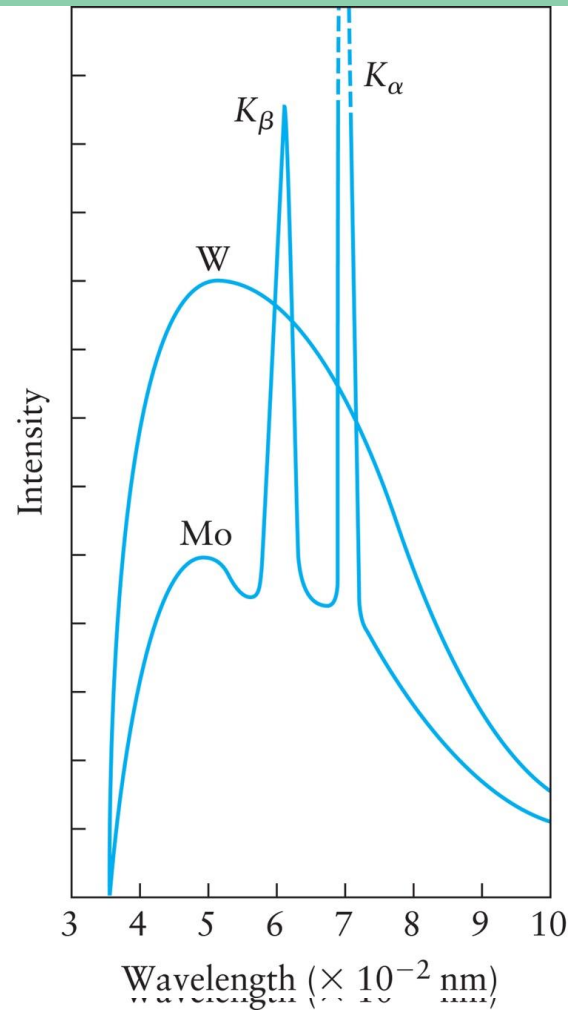
- After Pauli's discovery, it was determined that the elements in each column have similar ground-state configurations, and that is why they are alike.
- For example, hydrogen, lithium, and sodium are in the first column, called Group 1A, because each has one electron in its highest occupied energy level.
- This is one indication of how understanding quantum mechanics can play a crucial role in other scientific disciplines such as chemistry.

10.7 X-Ray Spectra

A band of ***X-rays*** are produced when electrons are accelerated to a high speed and then directed into a metal target, with different wavelengths having different intensities.

- The figure shows the x-ray spectra when the elements tungsten (W) and molybdenum (Mo) are used as targets.

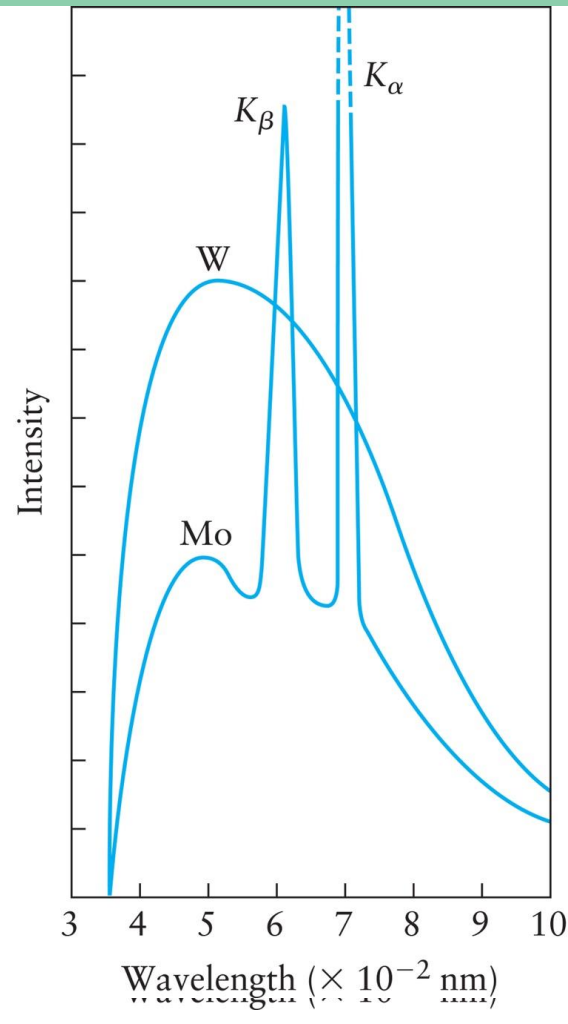
- Most of the x-rays are produced as the electrons are rapidly decelerated on entering the target.



10.7 X-Ray Spectra

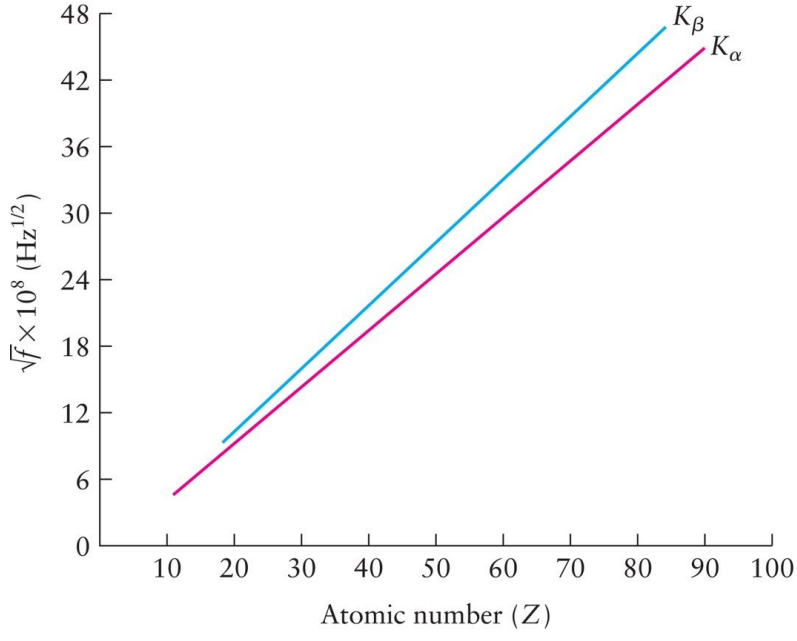
- The ***smooth part of the spectra*** covering the full range of wavelengths is called ***bremsstrahlung*** (German for "braking radiation").
- Bremsstrahlung spectra are much the same for different elements.

- But the ***two sharp peaks*** for the molybdenum target are unique to this element and constitute its ***characteristic spectrum***.
- Other elements exhibit characteristic spectra but at different wavelengths.



10.7 X-Ray Spectra

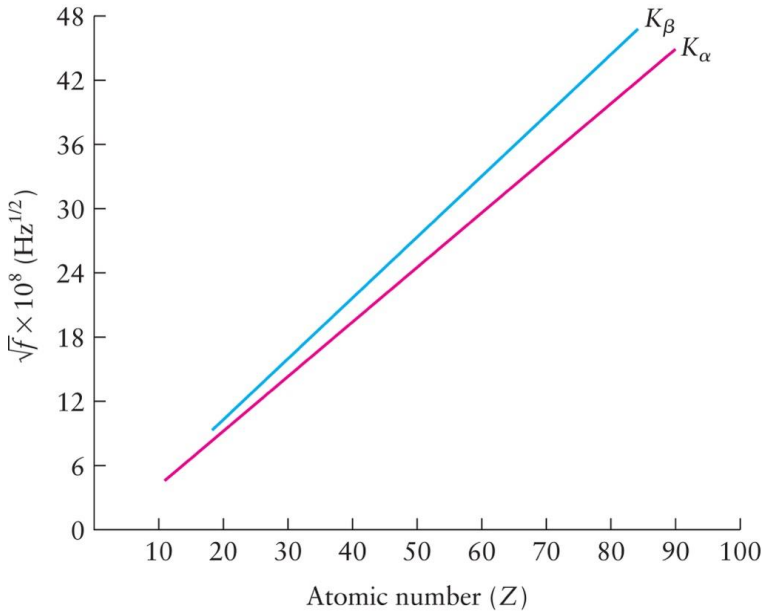
When the characteristic x-ray spectra of different elements was compared, it was found that there is a relationship between the frequencies of the peaks and the atomic number of the element.



- For each peak, a graph of the square root of the frequency versus the atomic number of the element was a straight line.

10.7 X-Ray Spectra

- This was a very practical discovery because it allowed the determination of the atomic number, and therefore the identity, of unknown elements.



- This was performed through determining the wavelengths of the characteristic x-ray peaks and then using this graph.

- This method was a key factor in the discovery of several new elements, including promethium (atomic number 61) and hafnium (atomic number 72).

10.7 X-Ray Spectra

- *It was proposed the following scenario to account for the x-ray spectra of the elements.*

- If a bombarding electron collides with, say, a molybdenum atom and knocks out an electron from the n = 1 level, electrons in the upper levels will cascade down to fill the vacancy left by the ejected electron, and photons will be emitted.

10.7 X-Ray Spectra

- The lowest-frequency (longest-wavelength) x-ray photon corresponds to the transition from $n = 2$ to $n = 1$. Because the lowest energy level was called the "**K shell**" by x-ray experimenters, this is called the ***K_α peak***.

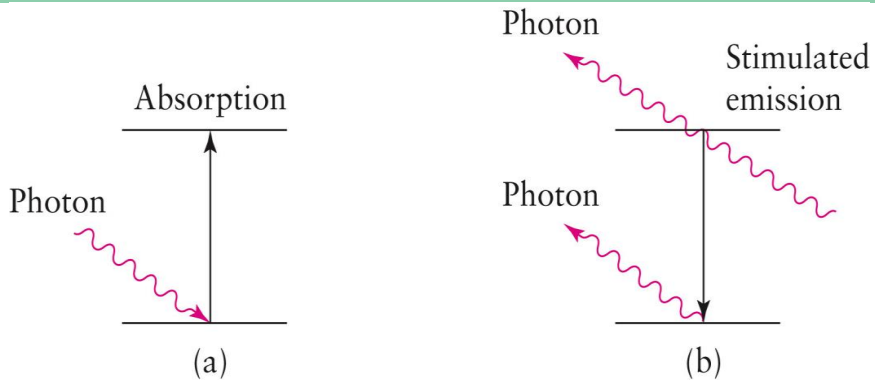
- Electrons jumping from $n = 3$ to $n = 1$ emit photons that form the ***K_β peak***.

- With this knowledge, we can see why high-speed electrons are needed to produce the characteristic x-rays of heavy elements. The inner electrons are so tightly bound that it takes very-high-energy electrons to knock them out.

10.8 Lasers

The word ***laser*** is an acronym derived from the phrase ***"light amplification by stimulated emission of radiation."***

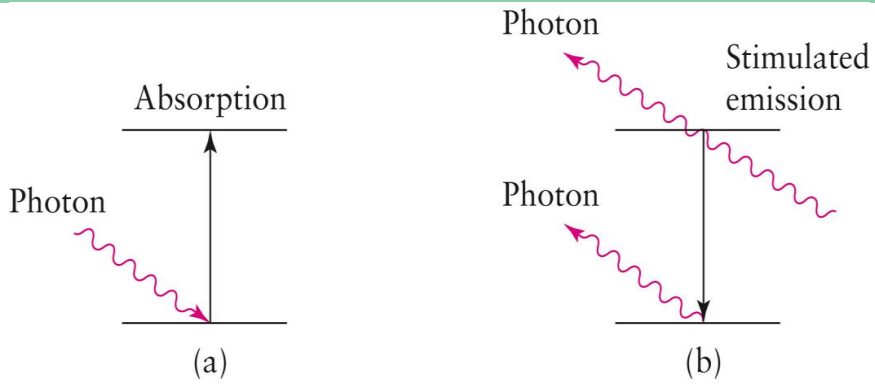
10.8 Lasers



- If an electron in an atom has been excited to some higher energy level by either a collision or the absorption of a photon with the proper energy, E.

- The electron can be stimulated to return to its original energy level through the intercession of a second photon with the same energy.
- If a group of atoms all having their electrons in this same excited state is "bathed" in light consisting of photons with energy E, the atoms will be stimulated to decay by emitting additional photons with the same energy E.

10.8 Lasers

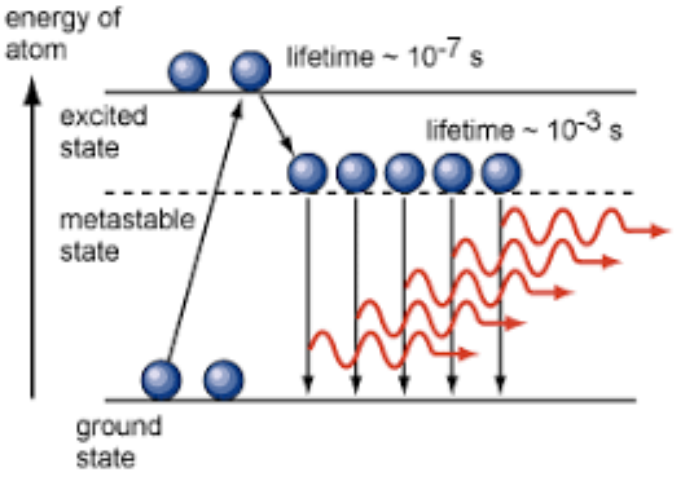


- By this process, the intensity of a beam of light is increased or amplified as a result of stimulated emission from atoms in the region through which the radiation passes. We have a **laser**.

- *To achieve this amplification*, we must first arrange for the majority of the atoms through which the stimulating radiation travels to have their electrons in the same excited state. Otherwise, the unexcited atoms will just absorb the radiation. This is not an easy situation to arrange because the excited atoms normally don't stay that way for long.

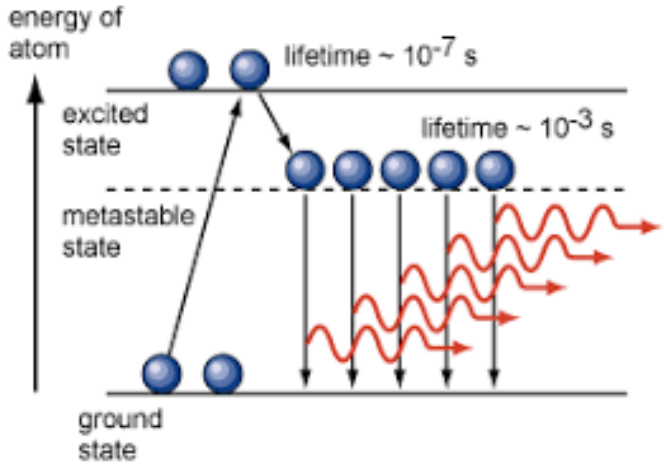
10.8 Lasers

The problem can be overcome because many atoms possess excited states that are said to be ***metastable***.



- Once in such a state, the electrons tend to remain there for a relatively long time (perhaps a thousandth of a second instead of a billionth) before spontaneously decaying to a lower state.
- During the time it takes to excite more than half of the atoms to that state, the ones already excited don't jump back down.

10.8 Lasers



- This process, called ***pumping***, can result in the majority of the atoms being in the same metastable, excited state.
- This condition is called a ***population inversion*** because there are more atoms with electrons populating an upper energy level than a lower one, the reverse of the usual situation. This condition is necessary for the generation of laser light.

10.8 Lasers

If we now irradiate the population-inverted atoms with photons of the correct energy, a chain reaction can be established that greatly amplifies the incident light beam.

- One photon stimulates the emission of another identical one, these two stimulate the emission of two more, these four stimulate the emission of four more, and so on.

- In the end, an avalanche of photons is produced, all having the same frequency (color).

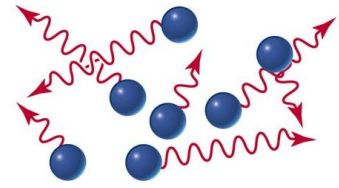
- The beam is said to be ***monochromatic***. The light amplification can produce a very high-intensity beam of light of a single wavelength.

10.8 Lasers

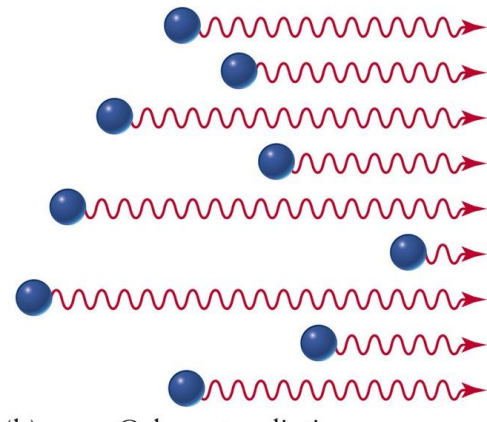
Besides being intense and monochromatic, laser light has an additional property that makes it extremely useful: ***coherence***.

- This means that the stimulating radiation and the additional emitted laser radiation are in phase: the crests and troughs of the EM waves at a given point all match up.

- An ordinary light source such as an incandescent bulb emits incoherent light (different parts of the beam are not in phase with one another).

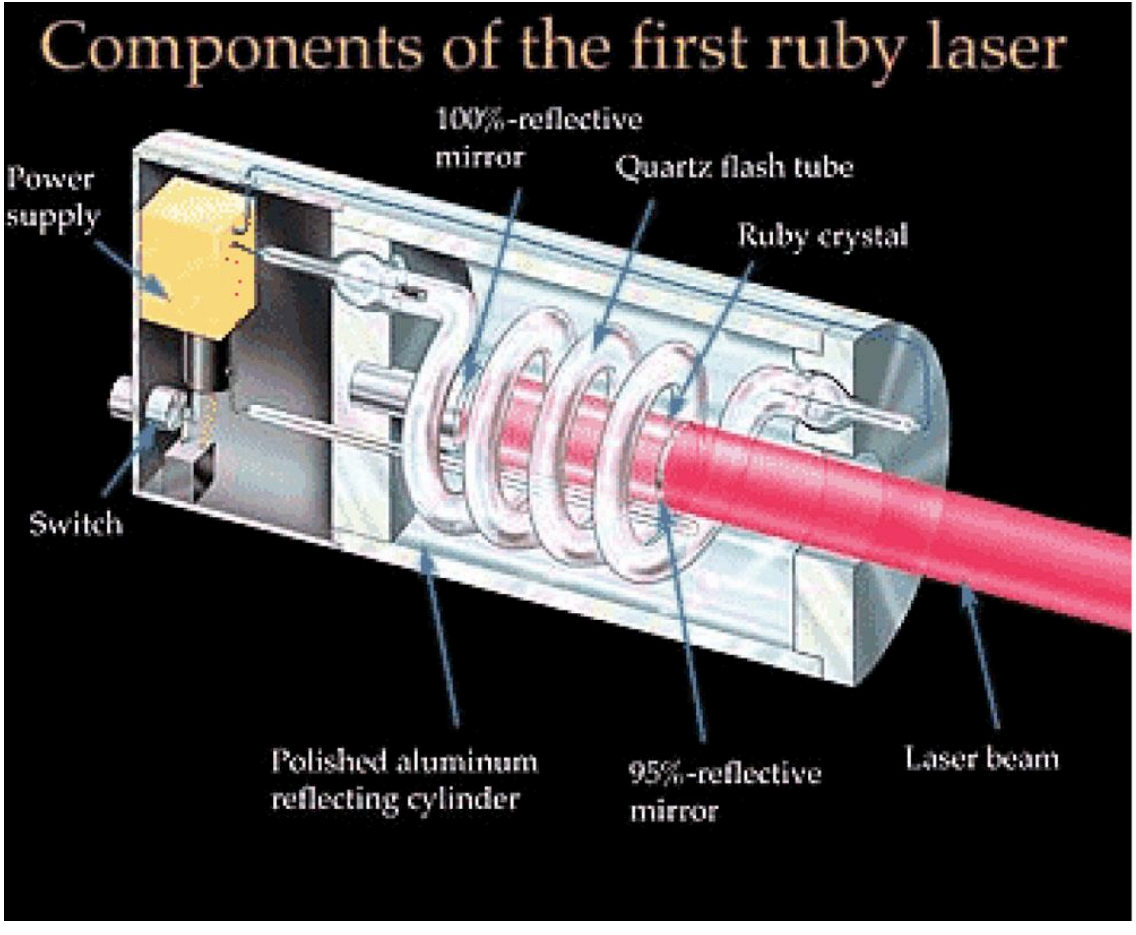


(a) Incoherent radiation from excited atoms



(b) Coherent radiation from excited atoms

10.8 Lasers


Components of the first ruby laser
100%-reflective mirror
Quartz flash tube
Power supply
Ruby crystal
Switch
Polished aluminum reflecting cylinder
95%-reflective mirror
Laser beam

Important Equations

IMPORTANT EQUATIONS

Equation	Comments
$E = hf$	Energy of atomic oscillator with frequency f. Energy of a photon with frequency f.
$\Delta E = hf$	Energy of photon emitted or absorbed in energy-level transition
allowed angular momentum $= n\left(\frac{h}{2\pi}\right)$	Quantization of electron orbital angular momentum (Bohr model)
$\lambda = \frac{h}{mv}$	de Broglie wavelength of material particle
$\Delta x \, \Delta mv \geq h$	Heisenberg's uncertainty principle
maximum number of allowed quantum states $= 2n^2$	Permitted number of electrons in each energy level (Pauli exclusion principle)