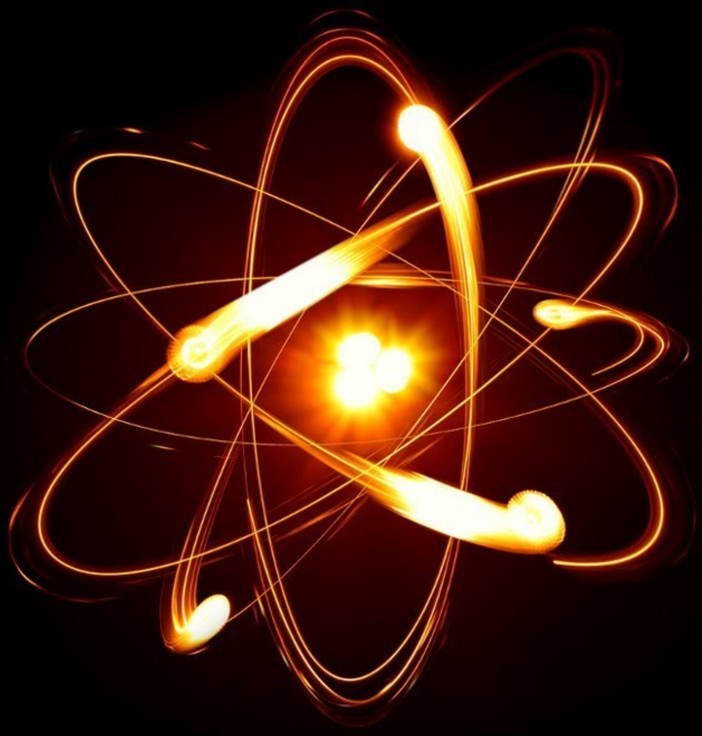
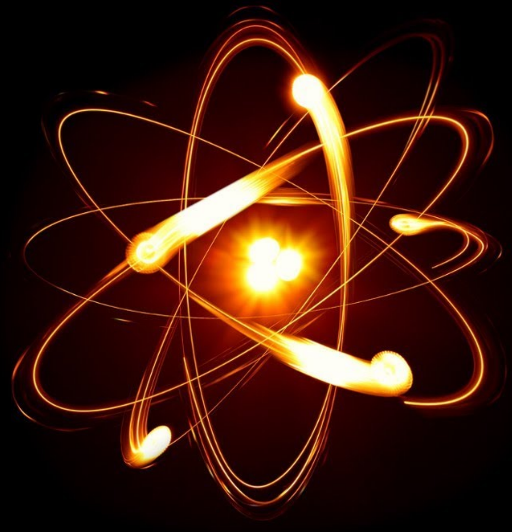
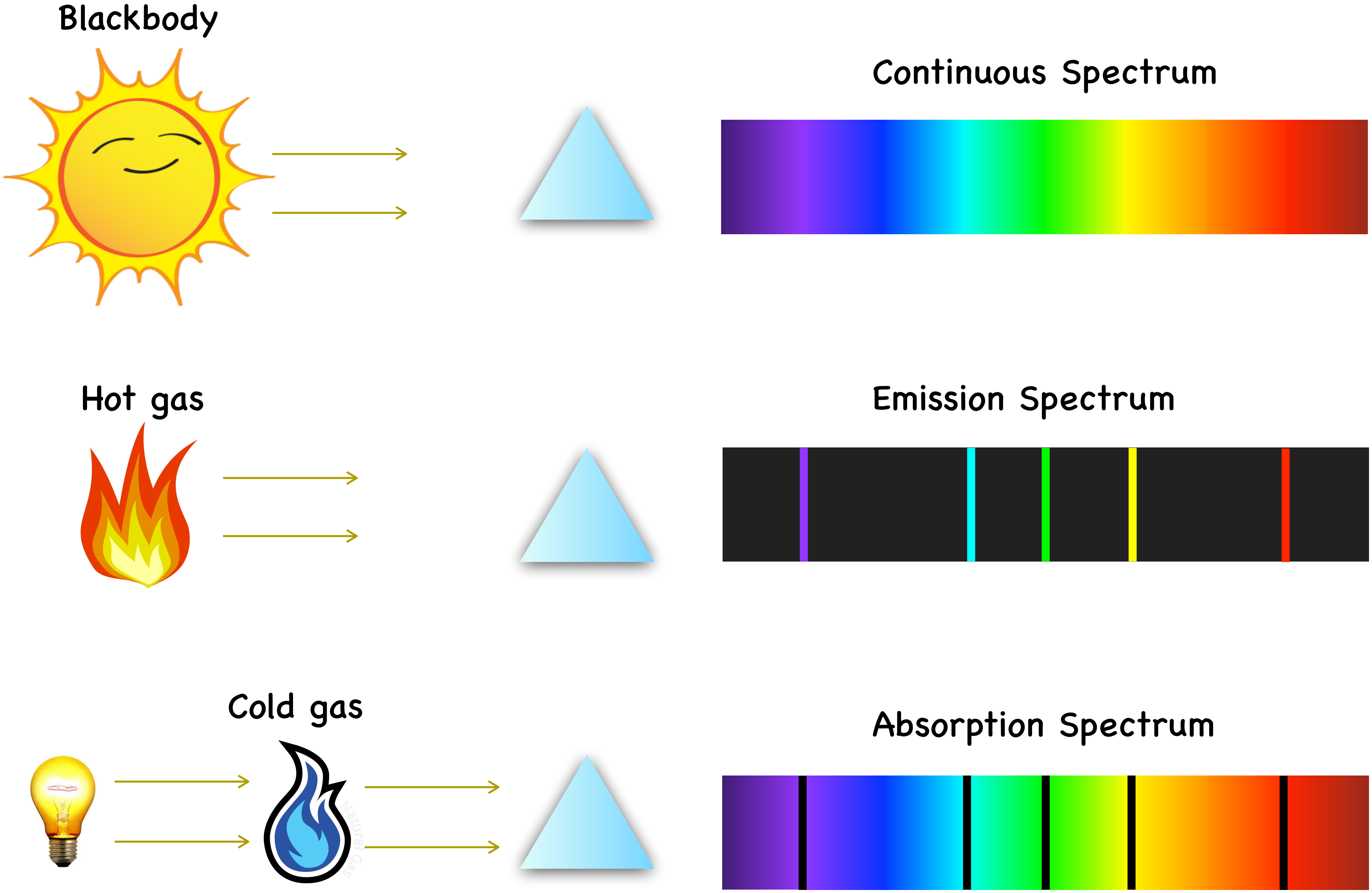
Ch.42: Atomic Physics



Chapter Outline





Blackbody

Continuous Spectrum

Hot gas

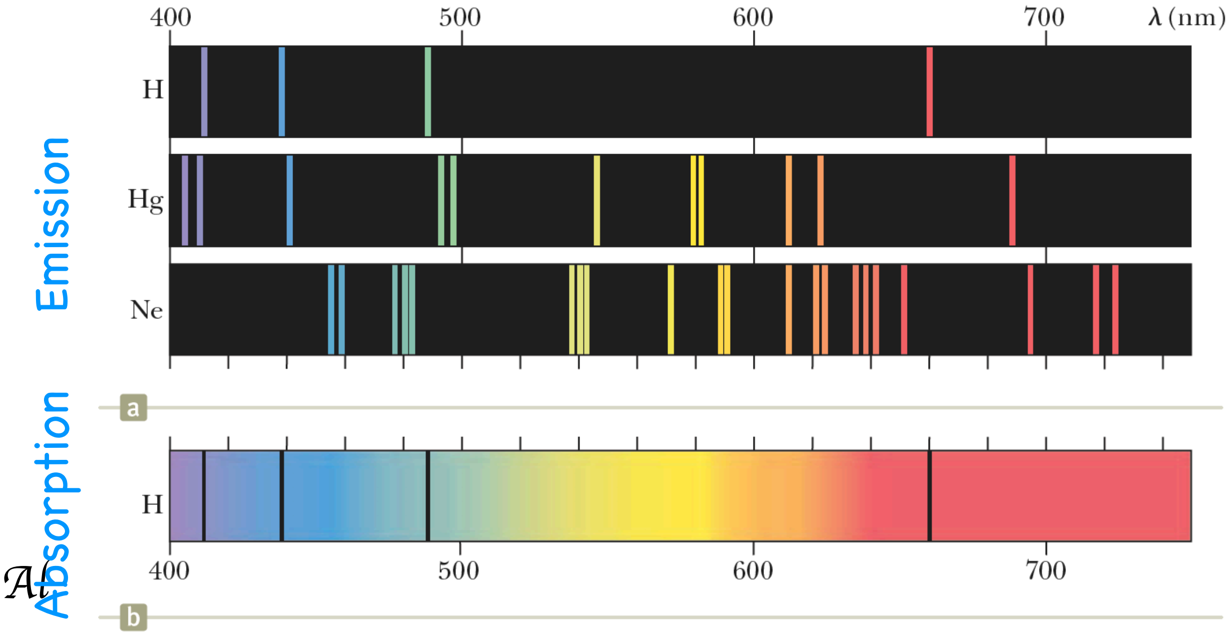
Emission Spectrum

Cold gas

Absorption Spectrum

42.1: Atomic Spectra of Gases

1. Heated objects emit thermal radiation that is characterized by *continuous distribution* of wavelengths (radiation with a range of wavelengths are emitted).
2. When a low-pressure gas undergoes an electric discharge, a *discrete line spectrum* is observed (only radiation with specific wavelengths are emitted, shown as colored lines). Observation and studying these spectral lines is called **emission spectroscopy**.
3. When a white light passes through a gas or dilute solution, a *discrete line spectrum* is observed, shown as black lines. Observation and studying these spectral lines is called **absorption spectroscopy**.




HAPPY
MOTHER'S
DAY


OPEN

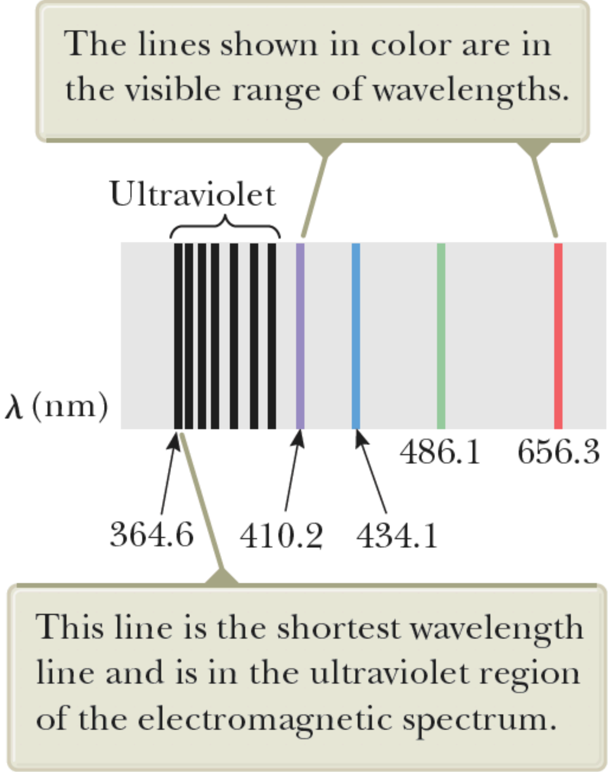
Balmer Series:

- Scientists tried to write down equations that can describe the atomic emission
- In 1885, Balmer found an empirical equation that correctly predicted the wavelengths of four visible emission lines of Hydrogen: H_α (red), H_β (Blue-green), H_γ (Blue-violet), H_δ (violet).

Balmer series ▶

$$\frac{1}{\lambda} = R_{\mathrm{H}}\left(\frac{1}{2^2} - \frac{1}{n^2}\right) \qquad n = 3, 4, 5, \ldots$$

- R_H is Rydberg constant = 1.0973732×10^7 m^{-1}
- n is an integer with values 3, 4, ,5 ,6 that give the four visible lines: $n = 3$ (red) and $n = 6$ (violet).
- The series limit is the shortest wavelength (364.6 nm) which corresponds to $n \to \infty$



The lines shown in color are in the visible range of wavelengths.

This line is the shortest wavelength line and is in the ultraviolet region of the electromagnetic spectrum.

Figure 42.2 The Balmer series of spectral lines for atomic hydrogen, with several lines marked with the wavelength in nanometers. (The horizontal wavelength axis is not to scale.)

Other Series:

$$\frac{1}{\lambda} = R_{\mathrm{H}}\left(1 - \frac{1}{n^2}\right) \qquad n = 2, 3, 4, \ldots \qquad \textbf{(42.2)}$$

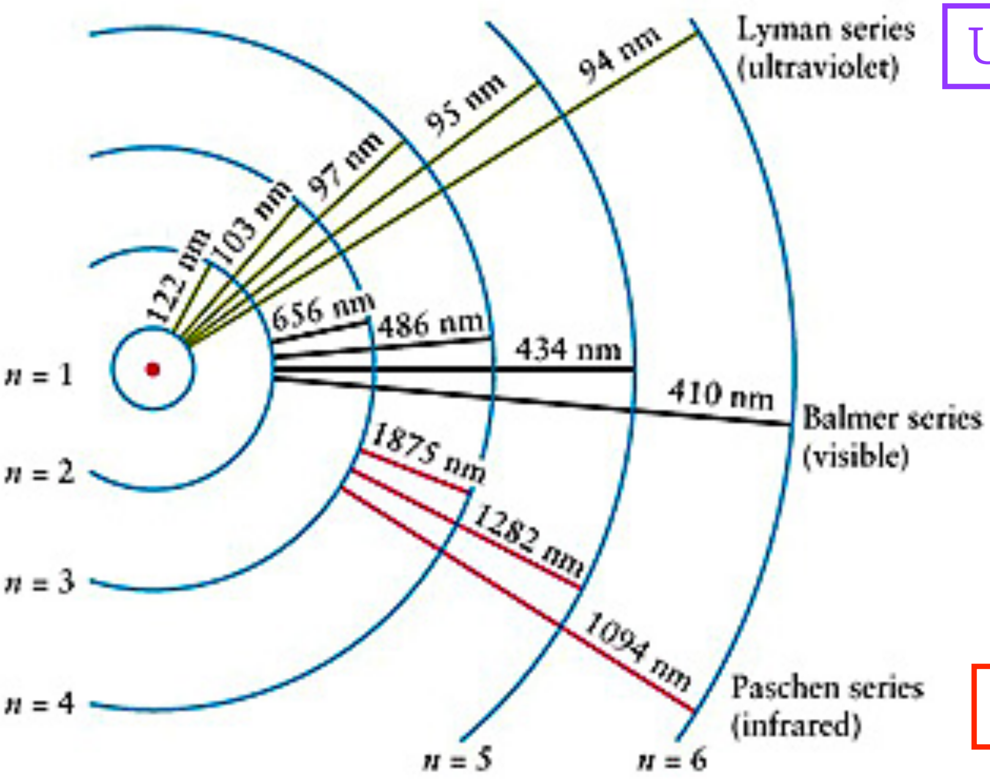
◀ **Lyman series**

$$\frac{1}{\lambda} = R_{\mathrm{H}}\left(\frac{1}{3^2} - \frac{1}{n^2}\right) \qquad n = 4, 5, 6, \ldots \qquad \textbf{(42.3)}$$

◀ **Paschen series**

$$\frac{1}{\lambda} = R_{\mathrm{H}}\left(\frac{1}{4^2} - \frac{1}{n^2}\right) \qquad n = 5, 6, 7, \ldots \qquad \textbf{(42.4)}$$

◀ **Brackett series**

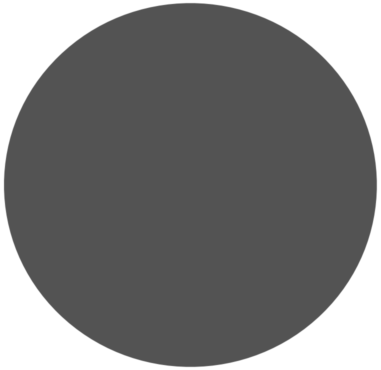


Ultraviolet

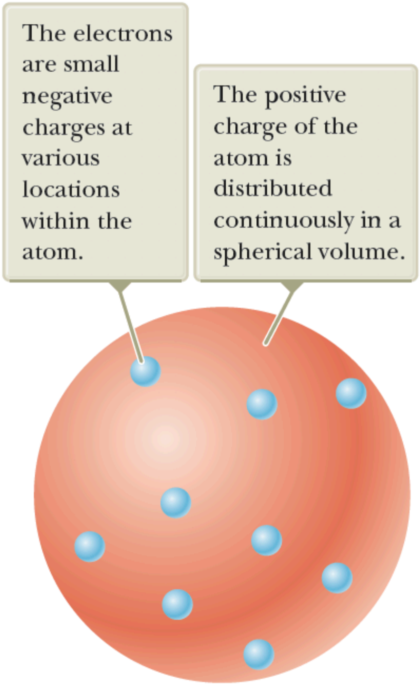
Visible

Infrared

42.2: Early Models of Atoms



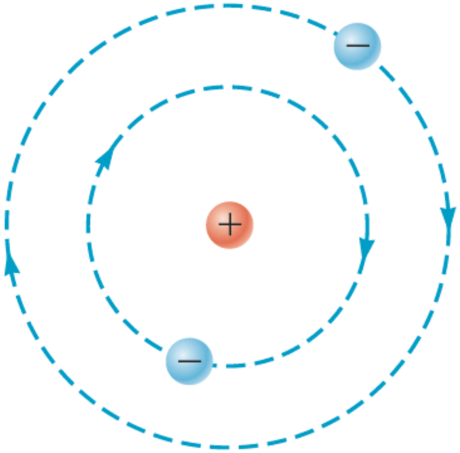
Atom in Newton Day's



The electrons are small negative charges at various locations within the atom.

The positive charge of the atom is distributed continuously in a spherical volume.

Thomson Model of Atom



Rutherford Planetary Model

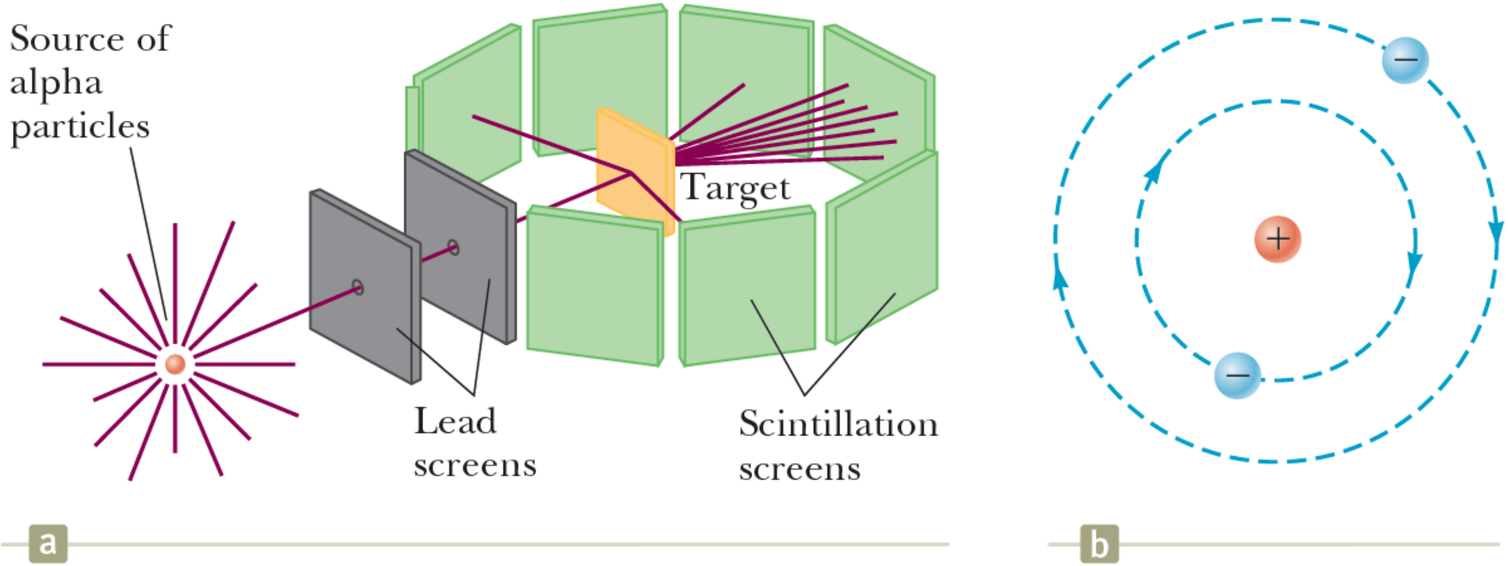


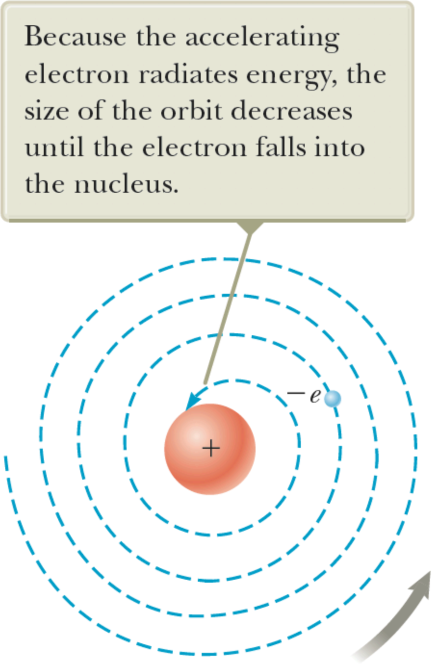
Figure 42.4 (a) Rutherford's technique for observing the scattering of alpha particles from a thin foil target. The source is a naturally occurring radioactive substance, such as radium. (b) Rutherford's planetary model of the atom.

Rutherford Atomic Model (Planetary Model):

1. The positive charge of the atom is concentrated in a region that was small relative to the size of the atom, and this concentrated positive charge is called the **nucleus** of the atom.
2. The atomic electrons are moving in orbits around the nucleus in the same manner as the planets orbit the sun.

Difficulties with Rutherford Model:

1. It does not explain the emission and absorption of radiation with specific frequencies by the atom.
2. When an electron moves in a circle around the nucleus, it will have a centripetal acceleration, and it should loose energy by emitting radiation. This means that the size of the orbit will decrease until the electron falls into the nucleus.



Because the accelerating electron radiates energy, the size of the orbit decreases until the electron falls into the nucleus.

Figure 42.5 The classical model of the nuclear atom predicts that the atom decays.

42.3: Bohr's Model of Hydrogen Atom

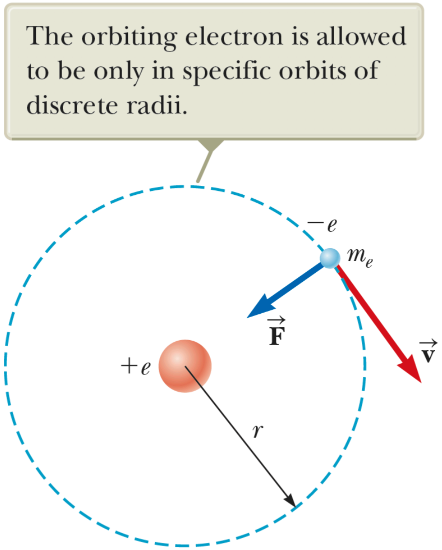

The orbiting electron is allowed to be only in specific orbits of discrete radii.

Figure 42.6 Diagram representing Bohr's model of the hydrogen atom.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Niels Bohr

Danish Physicist (1885–1962)

42.3: Bohr's Model of Hydrogen Atom

Bohr combined Plank's quantum theory, Einstein's concept of photons and Rutherford planetary model.

1. *Physical components:*
   The electron moves in circular orbits around the proton under the influence of the electric force of attraction as shown in Figure 42.6.

2. *Behavior of the components:*
   (a) Only certain electron orbits are stable. When in one of these **stationary states,** as Bohr called them, the electron does not emit energy in the form of radiation, even though it is accelerating. Hence, the total energy of the atom remains constant and classical mechanics can be used to describe the electron's motion. Bohr's model claims that the centripetally accelerated electron does not continuously emit radiation, losing energy and eventually spiraling into the nucleus, as predicted by classical physics in the form of Rutherford's planetary model.
   (b) The atom emits radiation when the electron makes a transition from a more energetic initial stationary state to a lower-energy stationary state. This transition cannot be visualized or treated classically. In particular, the frequency f of the photon emitted in the transition is related to the change in the atom's energy and is not equal to the frequency of the electron's orbital motion. The frequency of the emitted radiation is found from the energy-conservation expression

$$E_i - E_f = hf \qquad \textbf{(42.5)}$$

The orbiting electron is allowed to be only in specific orbits of discrete radii.

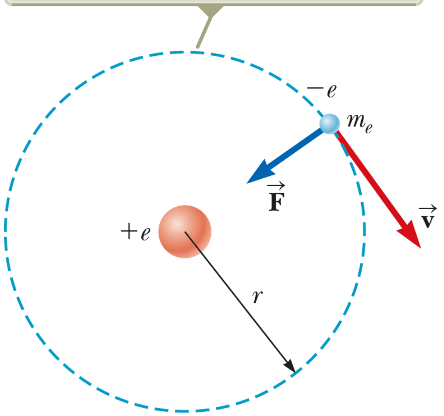

Figure 42.6 Diagram representing Bohr's model of the hydrogen atom.

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$$E_i - E_f = hf \tag{42.5}$$

where E_i is the energy of the initial state, E_f is the energy of the final state, and $E_i > E_f$. In addition, energy of an incident photon can be absorbed by the atom, but only if the photon has an energy that exactly matches the difference in energy between an allowed state of the atom and a higher-energy state. Upon absorption, the photon disappears and the atom makes a transition to the higher-energy state.

(c) The size of an allowed electron orbit is determined by a condition imposed on the electron's orbital angular momentum: the allowed orbits are those for which the electron's orbital angular momentum about the nucleus is quantized and equal to an integral multiple of $\hbar = h/2\pi$,

$$m_e vr = n\hbar \quad n = 1, 2, 3, \ldots \tag{42.6}$$

where m_e is the electron mass, v is the electron's speed in its orbit, and r is the orbital radius.

Velocity of electrons around the nucleus - for H:

$$F_c = F_e$$

$$ma_c = m\frac{v^2}{r} = k\frac{e^2}{r^2}$$

$$v = \sqrt{k\frac{e^2}{mr}}$$

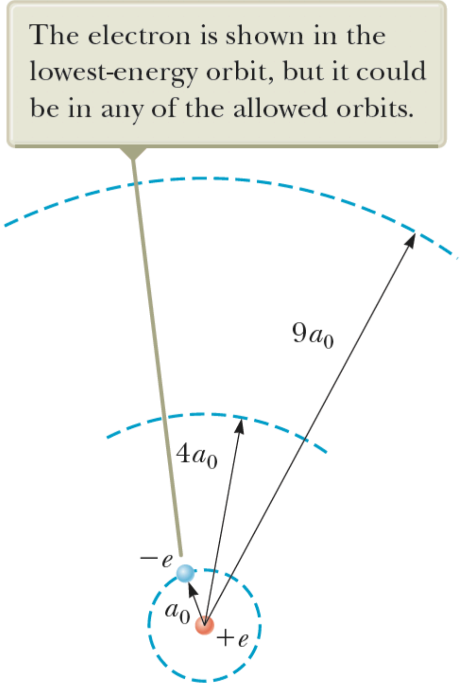
The electron is shown in the lowest-energy orbit, but it could be in any of the allowed orbits.

Figure 42.7 The first three circular orbits predicted by the Bohr model of the hydrogen atom.

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$$mvr = n\hbar$$

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

Bohr's radius (a_0):

$$a_0 = \frac{\hbar^2}{m_e k_e e^2} = 0.052\,9\ \text{nm}$$

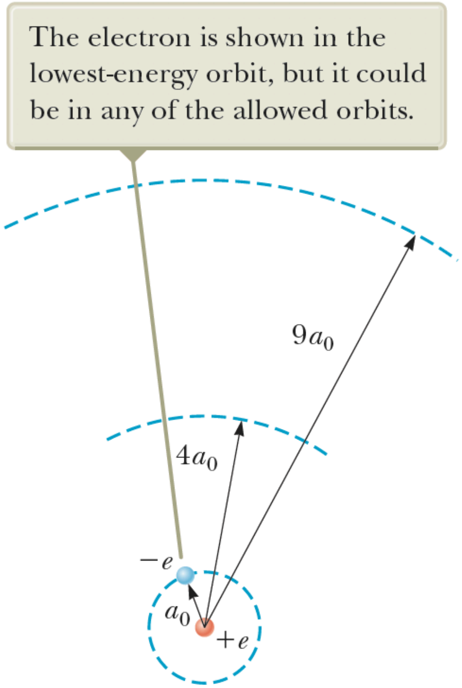
Radii of Bohr orbits in hydrogen:

$$r_n = n^2 a_0 = n^2(0.052\,9\ \text{nm}) \qquad n = 1, 2, 3, \ldots$$

Allowed energy levels of Hydrogen:

$$E_n = -\frac{k_e e^2}{2a_0}\left(\frac{1}{n^2}\right) \qquad n = 1, 2, 3, \ldots$$

$$E_n = -\frac{13.606\ \text{eV}}{n^2} \qquad n = 1, 2, 3, \ldots$$



The electron is shown in the lowest-energy orbit, but it could be in any of the allowed orbits.

Figure 42.7 The first three circular orbits predicted by the Bohr model of the hydrogen atom.

- If the energy of the atom is raised from that of the ground state to any energy larger than zero, the atom is said to be *ionized*.
- **Ionization Energy**: the minimum energy required to ionize the atom in its ground state.
- From Fig.42.8: The ionization energy for Hydrogen is 13.6 eV

The colored arrows for the Balmer series indicate that this series results in the emission of visible light.

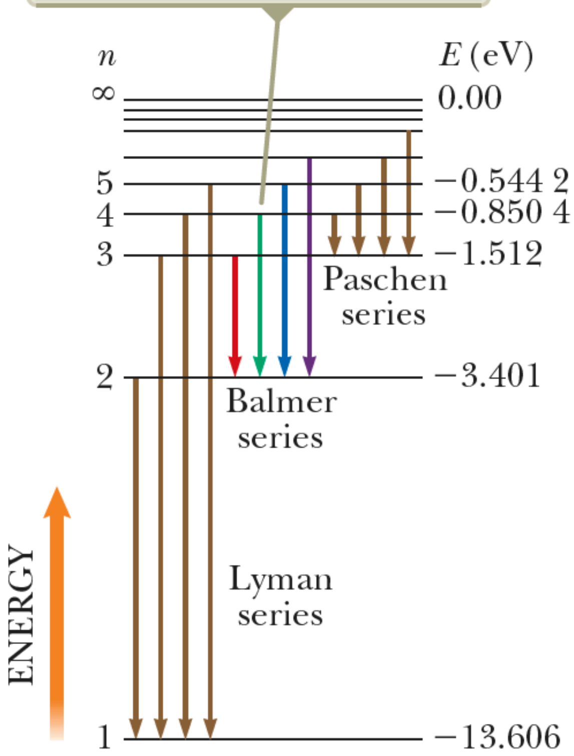


Figure 42.8 An energy-level diagram for the hydrogen atom. Quantum numbers are given on the left, and energies (in electron volts) are given on the right. Vertical arrows represent the four lowest-energy transitions for each of the spectral series shown.

Frequency of photon emitted when an electron makes a transition from an outer orbit to an inner orbit:

$$f = \frac{E_i - E_f}{h} = \frac{k_e e^2}{2a_0 h}\left(\frac{1}{n_f^{\,2}} - \frac{1}{n_i^{\,2}}\right)$$

The wavelength of the photon:

$$\frac{1}{\lambda} = \frac{f}{c} = \frac{k_e e^2}{2a_0 hc}\left(\frac{1}{n_f^{\,2}} - \frac{1}{n_i^{\,2}}\right)$$

$$\boxed{\frac{1}{\lambda} = R_{\mathrm{H}}\left(\frac{1}{n_f^{\,2}} - \frac{1}{n_i^{\,2}}\right)}$$

For atoms other than hydrogen (Z>1):

$$r_n = (n^2)\frac{a_0}{Z}$$

$$E_n = -\frac{k_e e^2}{2a_0}\left(\frac{Z^2}{n^2}\right) \quad n = 1, 2, 3, \ldots$$

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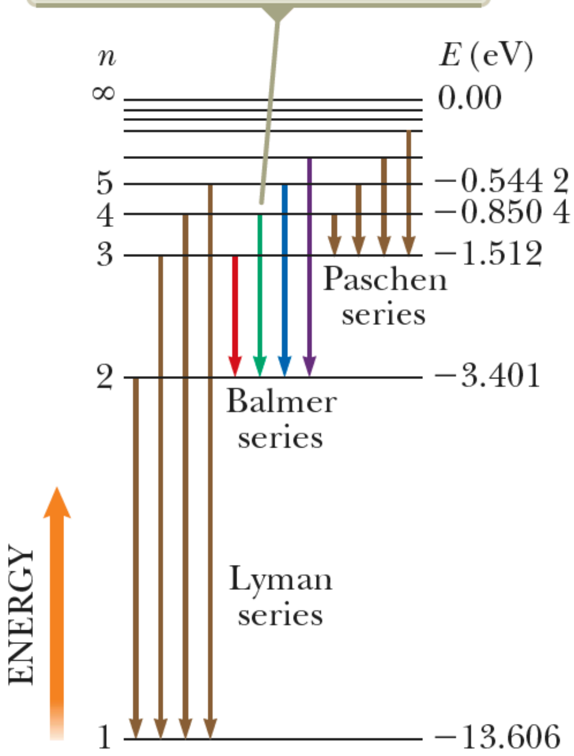
The colored arrows for the Balmer series indicate that this series results in the emission of visible light.



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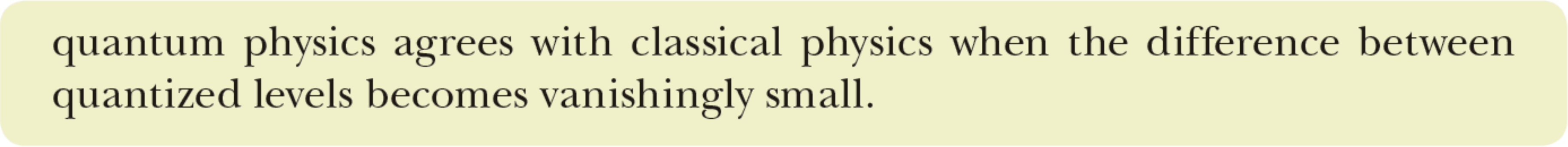
Bohr Correspondance Principle:



quantum physics agrees with classical physics when the difference between quantized levels becomes vanishingly small.

Example 42.1 Electronic Transitions in Hydrogen

(A) The electron in a hydrogen atom makes a transition from the $n = 2$ energy level to the ground level ($n = 1$). Find the wavelength and frequency of the emitted photon.

(B) In interstellar space, highly excited hydrogen atoms called Rydberg atoms have been observed. Find the wavelength to which radio astronomers must tune to detect signals from electrons dropping from the $n = 273$ level to the $n = 272$ level.

(C) What is the radius of the electron orbit for a Rydberg atom for which $n = 273$?

(D) How fast is the electron moving in a Rydberg atom for which $n = 273$?

42.8: More on Atomic Spectra: Visible and X-Ray

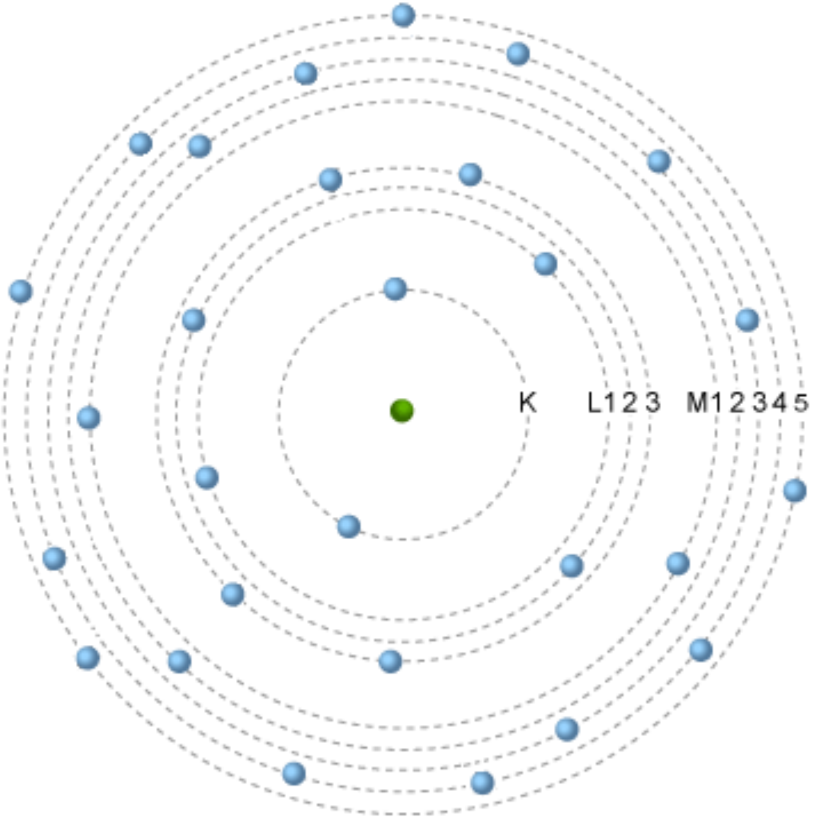
- According to Bohr, the allowed energy levels of on-electron atoms and ions such as hydrogen and He^+

$$E_n = -\frac{k_e e^2}{2a_0}\left(\frac{Z^2}{n^2}\right) = -\frac{(13.6\text{ eV})Z^2}{n^2}$$

- For multi-electron atoms, the nuclear positive charge Ze is shielded by the inner-shell electrons. Therefore, for multi-electron atoms the atomic number Z is replaced by the effective atomic number Z_{eff} to take into account the positive charge that the electron in an outer shell is *really seeing*

$$E_n = -\frac{(13.6\text{ eV})Z_{\text{eff}}^2}{n^2}$$

42.8: More on Atomic Spectra: Visible and X-Ray



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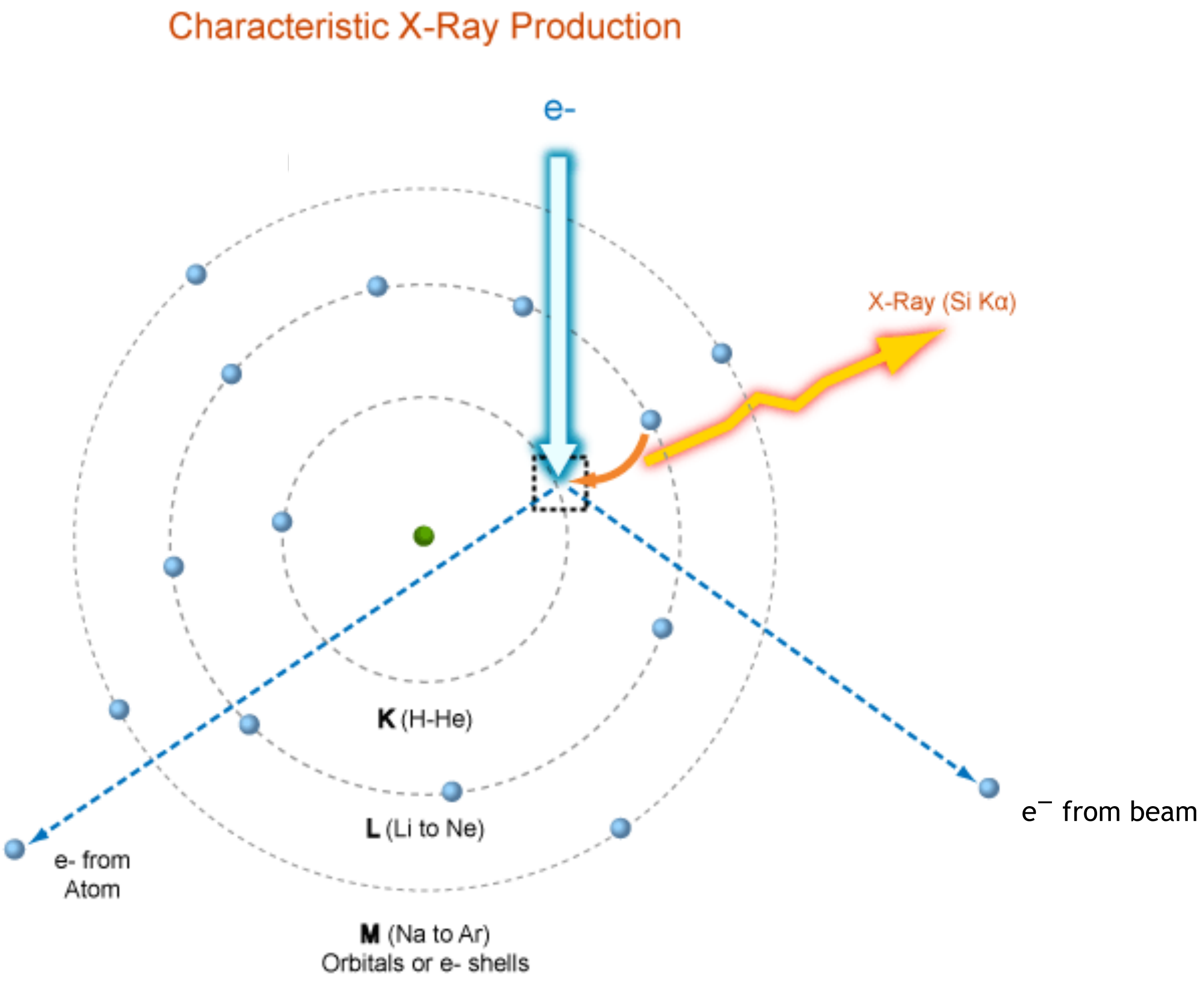

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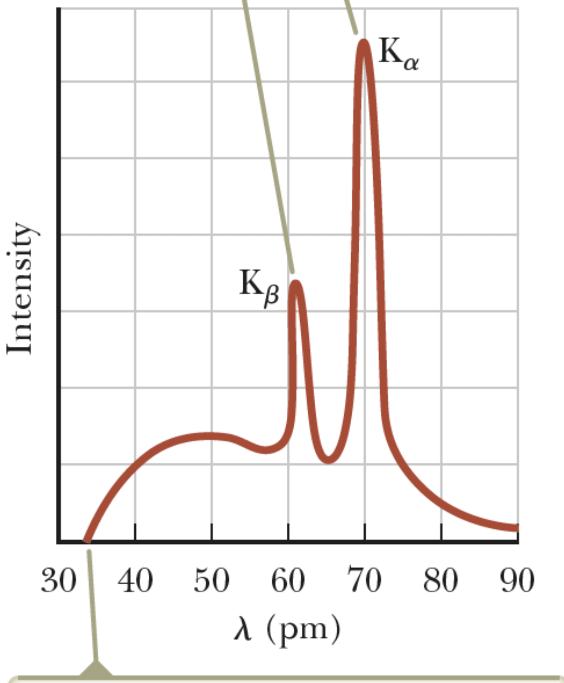
42.8: More on Atomic Spectra: Visible and X-Ray

- X-rays are commonly produced by accelerating (or decelerating) charged particles. When a fast electron decelerate, it emits radiation called *bremsstrahlung*.
- Examples include a beam of electrons striking a metal plate.
- Next Figure, shows X-ray spectra. What do you see?
- X-rays are emitted with discrete wavelengths characteristic of the energy level spacings in the atoms
- *Characteristic X-rays* are produced by electron transitions between the electron shells.
- The peaks represent characteristic X-rays. They are formed when electrons drop from upper level to vacancies other than those in the K shell.
- K_α is produced when a vacancy in K shell is filed by an electron from the next higher level (L), and so on.

The peaks represent *characteristic x-rays*. Their appearance depends on the target material.



The continuous curve represents *bremsstrahlung*. The shortest wavelength depends on the accelerating voltage.

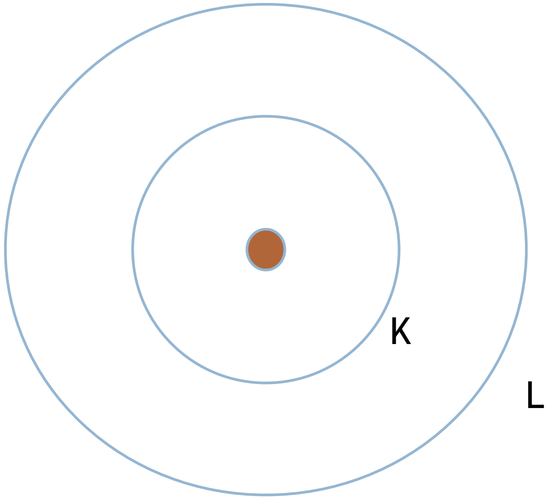
X-Ray Spectra:

- Consider an atom with atomic number Z in which one of the two electrons in the K shell has been ejected.
- Since the electron in the L shell will only see an effective charge (Z-1), the energy of an electron in the L shell:

$$E_{\mathrm{L}} = -(Z-1)^2 \frac{13.6\ \mathrm{eV}}{2^2}$$

After the atom makes the transition, there are two electrons in the K shell. We can estimate the energy associated with on these electrons as:

$$E_{\mathrm{K}} \approx -Z^2(13.6\ \mathrm{eV})$$

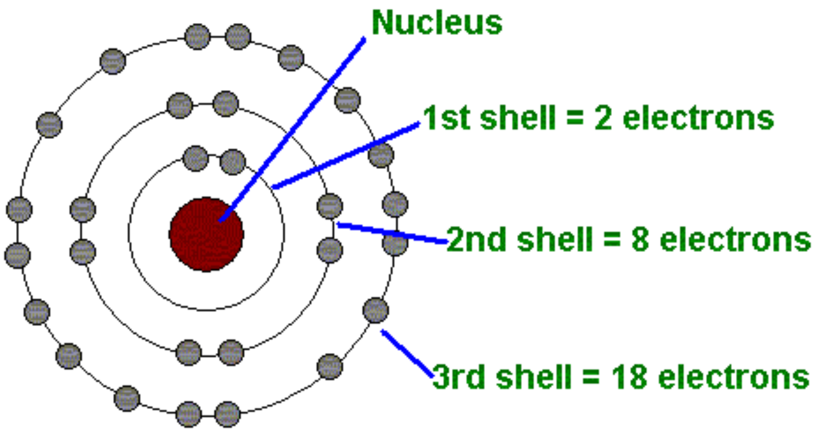


- The energy of the emitted X-ray due to the transition of the electron from L to K shell is:

$$E_{X-ray} = hf = E_L - E_K$$

Example 42.5 Estimating the Energy of an X-Ray

Estimate the energy of the characteristic x-ray emitted from a tungsten target when an electron drops from an M shell ($n = 3$ state) to a vacancy in the K shell ($n = 1$ state). The atomic number for tungsten is $Z = 74$.



42.9: Spontaneous and Stimulated Transitions

- When radiation is incident on the atom, only those photons whose energy *hf* matches the energy separation ΔE between two energy levels can be absorbed by the atom. This process is called **Stimulated absorption**.
- Once an atom is in an excited state, the excited atom can make a transition back to a lower energy level, emitting a photon in a process called **Spontaneous emission.**

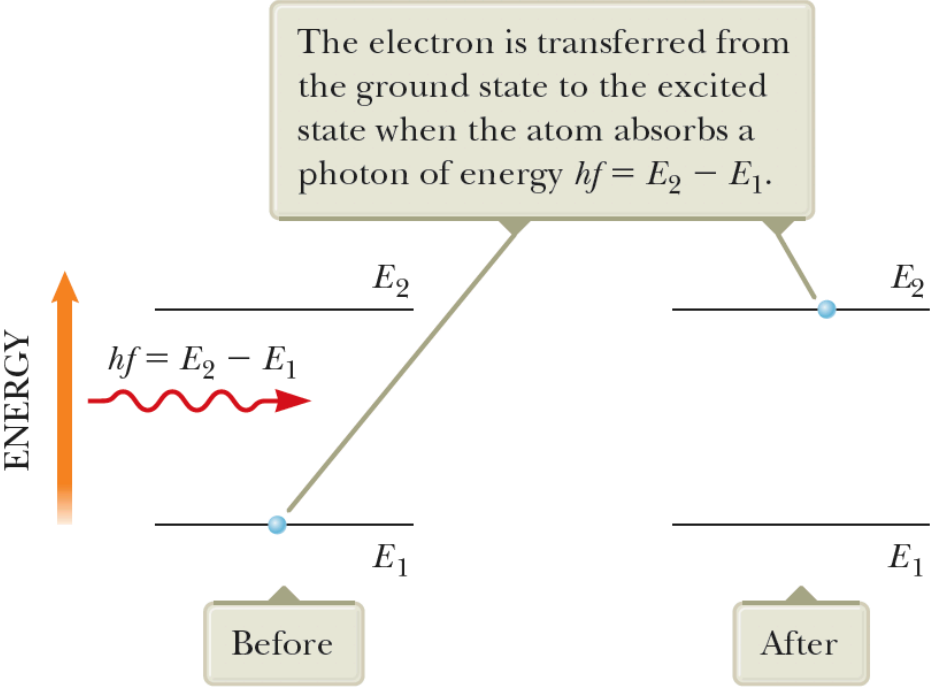

The electron is transferred from the ground state to the excited state when the atom absorbs a photon of energy $hf = E_2 - E_1$.

Figure 42.26 Stimulated absorption of a photon.

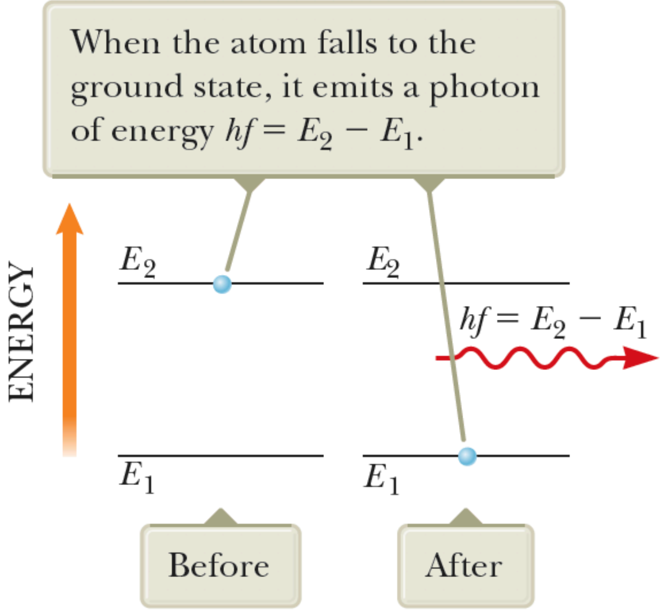

When the atom falls to the ground state, it emits a photon of energy $hf = E_2 - E_1$.

Figure 42.27 Spontaneous emission of a photon by an atom that is initially in the excited state E_2.

42.9: Spontaneous and Stimulated Transitions

- Metastable state: is the excited state of the atom that has a relatively large lifetime (larger than 10^{-8} s).
- When photon of energy $E_2 - E_1$ is incident on the atom. One possibility is that the photon energy is sufficient for the photon to ionize the atom. Another possibility is that the interaction between the incoming photon and the atom causes the atom to *return to the ground state* and emit a second photon.

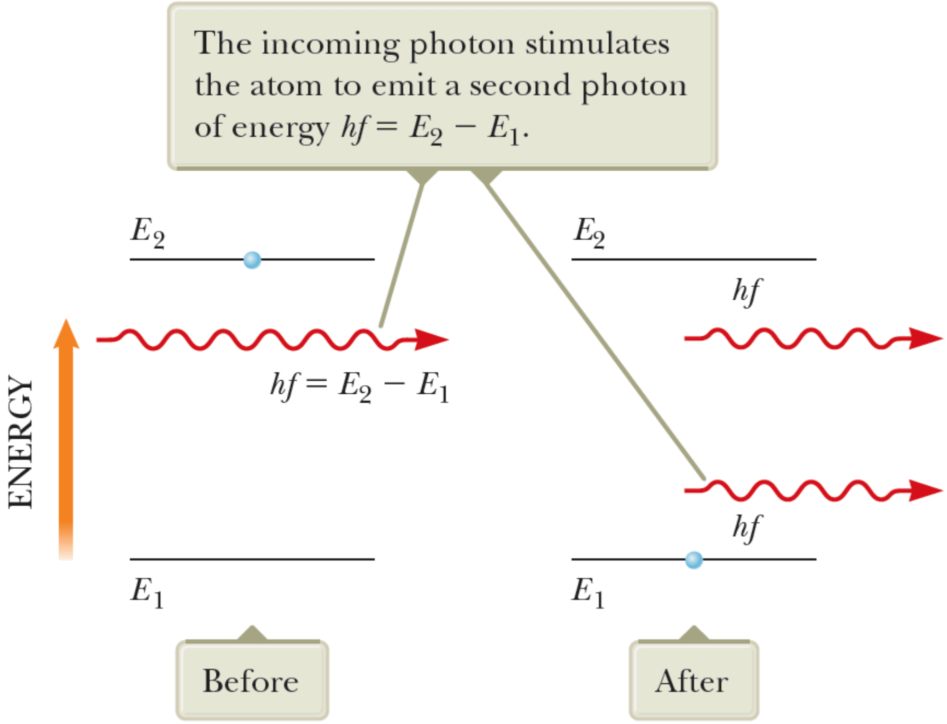


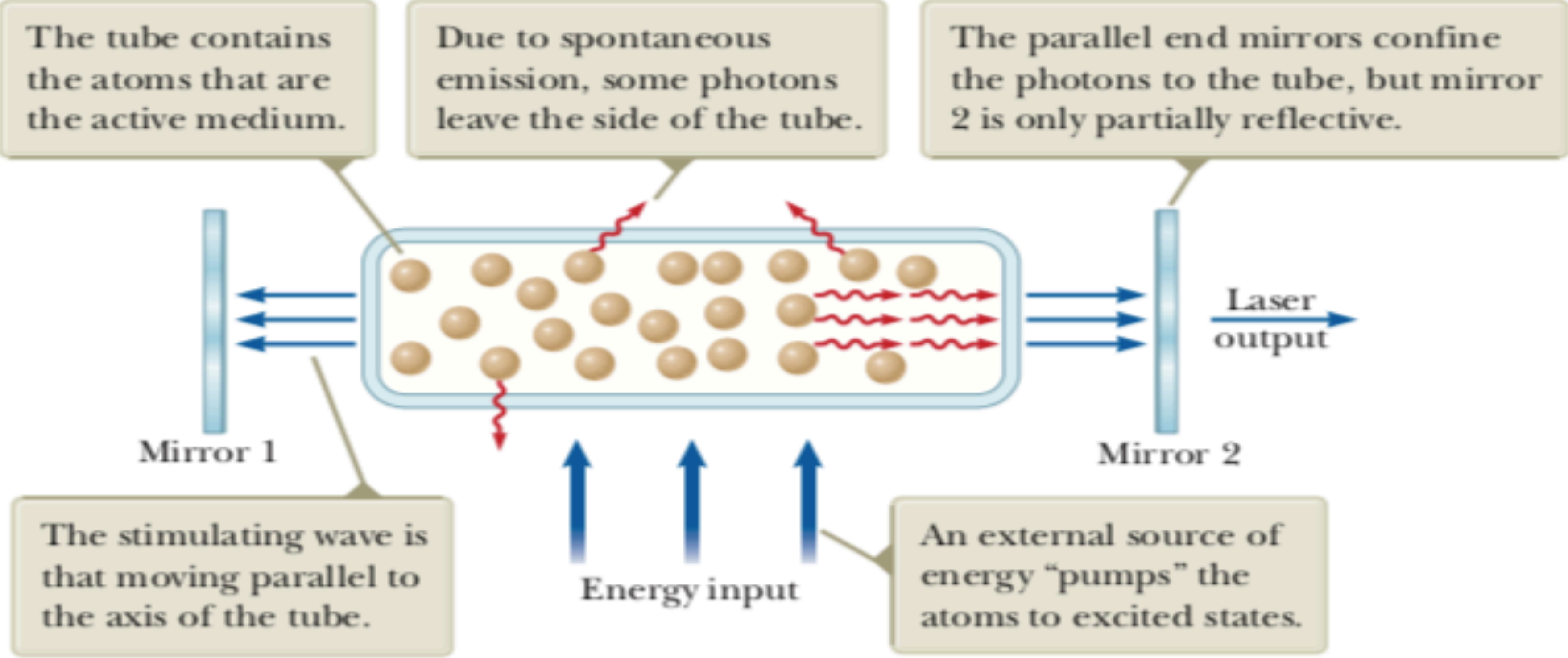
Figure 42.28 Stimulated emission of a photon by an incoming photon of energy $hf = E_2 - E_1$. Initially, the atom is in the excited state.

42.10: Lasers (Light Amplification by Simulated Emission of Radiation)

- Laser light is coherent. The individual rays of light in a laser beam maintain a fixed phase relationship with one another.
- Laser light is monochromatic. Light in a laser beam has a very narrow range of wavelengths.
- Laser light has a small angle of divergence. The beam spreads out very little, even over large distances.

- **Populated Inversion**: is the condition where more atoms of the system are in excited states than in the ground state.

- The system must be in a state of population inversion: there must be more atoms in an excited state than in the ground state. That must be true because the number of photons emitted must be greater than the number absorbed.
- The excited state of the system must be a *metastable state,* meaning that its lifetime must be long compared with the usually short lifetimes of excited states, which are typically 10^{-8} s. In this case, the population inversion can be established and stimulated emission is likely to occur before spontaneous emission.
- The emitted photons must be confined in the system long enough to enable them to stimulate further emission from other excited atoms. That is achieved by using reflecting mirrors at the ends of the system. One end is made totally reflecting, and the other is partially reflecting. A fraction of the light intensity passes through the partially reflecting end, forming the beam of laser light (Fig. 42.29).



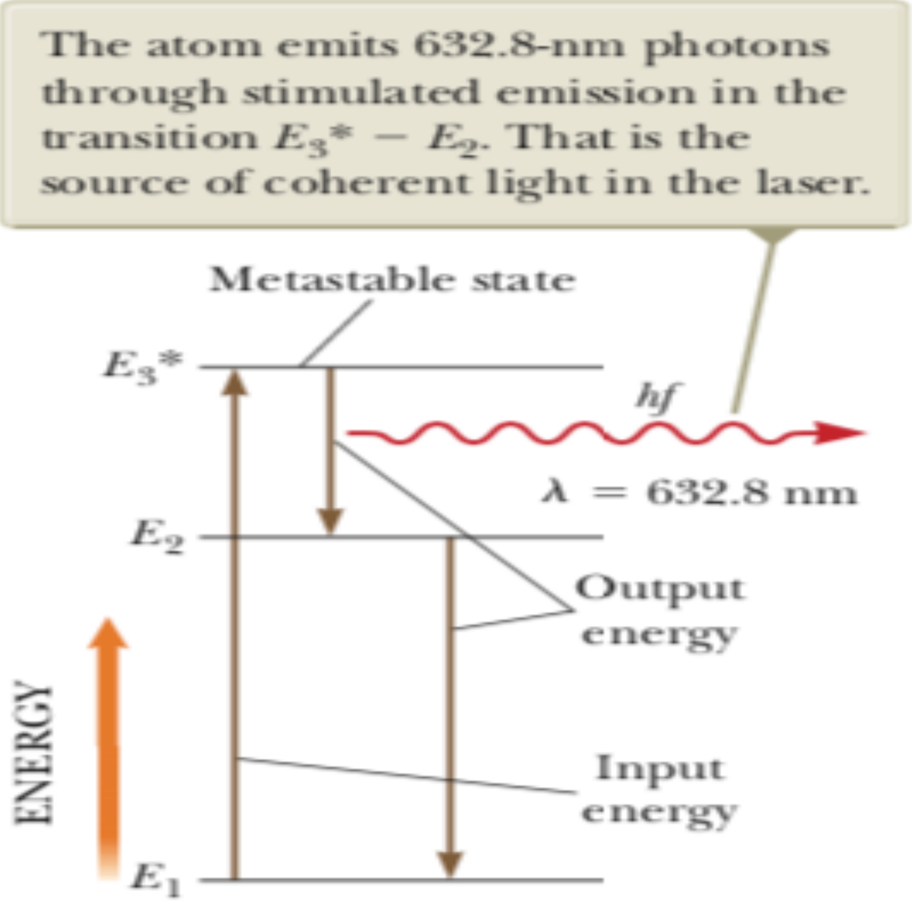
The tube contains the atoms that are the active medium.

Due to spontaneous emission, some photons leave the side of the tube.

The parallel end mirrors confine the photons to the tube, but mirror 2 is only partially reflective.

The stimulating wave is that moving parallel to the axis of the tube.

An external source of energy “pumps” the atoms to excited states.



The atom emits 632.8-nm photons through stimulated emission in the transition $E_3{}^* - E_2$. That is the source of coherent light in the laser.

Figure 42.30 Energy-level diagram for a neon atom in a helium–neon laser.