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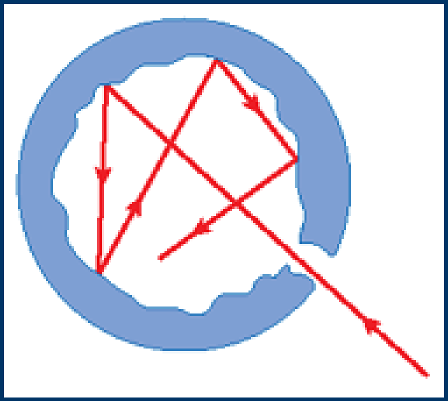
Blackbody Radiation

- Physicists study **blackbody radiation** by observing a hollow object with a small opening, as shown in the diagram.
- A **blackbody** is a perfect radiator and absorber and emits radiation based only on its temperature.



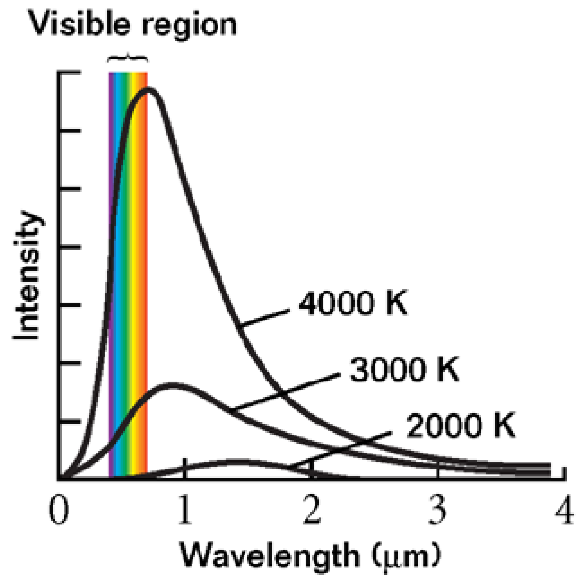
Light enters this hollow object through the small opening and strikes the interior wall. Some of the energy is absorbed, but some is reflected at a random angle. After many reflections, essentially all of the incoming energy is absorbed by the cavity wall.

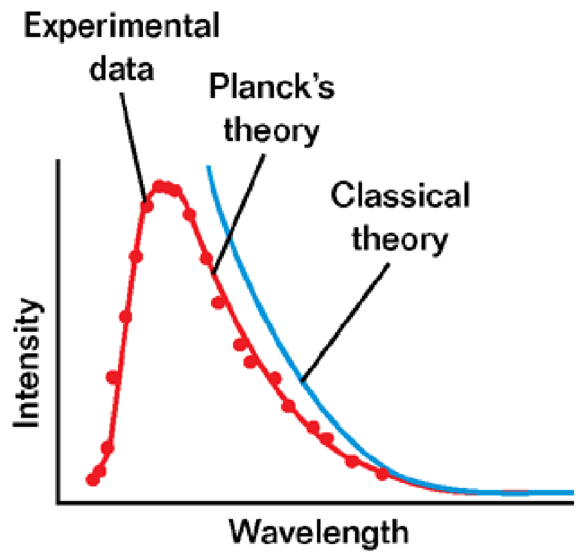
Blackbody Radiation, *continued*

- The **ultraviolet catastrophe** is the failed prediction of classical physics that the energy radiated by a blackbody at extremely short wavelengths is extremely large and that the total energy radiated is infinite.

- **Max Planck** (1858–1947) developed a formula for blackbody radiation that was in complete agreement with experimental data at all wavelengths by assuming that **energy comes in discrete units,** or is **quantized.**

Blackbody Radiation





The graph on the left shows the intensity of blackbody radiation at three different temperatures. Classical theory's prediction for blackbody radiation (the blue curve) did not correspond to the experimental data (the red data points) at all wavelengths, whereas Planck's theory (the red curve) did.

Quantum Energy

- Einstein later applied the concept of quantized energy to light. The units of light energy called **quanta** (now called **photons**) are absorbed or given off as a result of electrons "jumping" from one quantum state to another.
- The **energy of a light quantum,** which corresponds to the energy difference between two adjacent levels, is given by the following equation:

$$E = hf$$

energy of a quantum = Planck's constant × frequency

Planck's constant (h) ≈ 6.63×10^{-34} J•s

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Quantum Energy

- If Planck's constant is expressed in units of J•s, the equation $E = hf$ gives the energy in **joules.**
- However, in atomic physics, energy is often expressed in units of the **electron volt, eV.**
- An **electron volt** is defined as the energy that an electron or proton gains when it is accelerated through a potential difference of 1 V.
- The relation between the electron volt and the joule is as follows:

$$1 \text{ eV} = 1.60 \times 10^{-19} \text{ J}$$

The Photoelectric Effect

- The **photoelectric effect** is the emission of electrons from a material surface that occurs when light of certain frequencies shines on the surface of the material.
- Classical physics cannot explain the photoelectric effect.
- Einstein assumed that an electromagnetic wave can be viewed as a stream of particles called **photons.** Photon theory accounts for observations of the photoelectric effect.

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The Photoelectric Effect

	Classical predictions	Experimental evidence
Whether electrons are ejected depends on . . .	the intensity of the light.	the frequency of the light.
The kinetic energy of ejected electrons depends on . . .	the intensity of the light.	the frequency of the light.
At low intensities, electron ejection . . .	takes time.	occurs almost instantaneously above a certain frequency.

The Photoelectric Effect, *continued*

- No electrons are emitted if the frequency of the incoming light falls below a certain frequency, called the **threshold frequency (f_t)**.
- The smallest amount of energy the electron must have to escape the surface of a metal is the **work function** of the metal.
- The work function is equal to hf_t.

The Photoelectric Effect, *continued*

Because energy must be conserved, the maximum kinetic energy (of photoelectrons ejected from the surface) is the difference between the photon energy and the work function of the metal.

maximum kinetic energy of a photoelectron

$$KE_{max} = hf - hf_t$$

maximum kinetic energy = (Planck's constant × frequency of incoming photon) – work function

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Compton Shift

- If light behaves like a particle, then **photons should have momentum** as well as energy; both quantities should be conserved in elastic collisions.
- The American physicist **Arthur Compton** directed X rays toward a block of graphite to test this theory.
- He found that the scattered waves had **less energy** and **longer wavelengths** than the incoming waves, just as he had predicted.
- This change in wavelength, known as the **Compton shift,** supports Einstein's photon theory of light.

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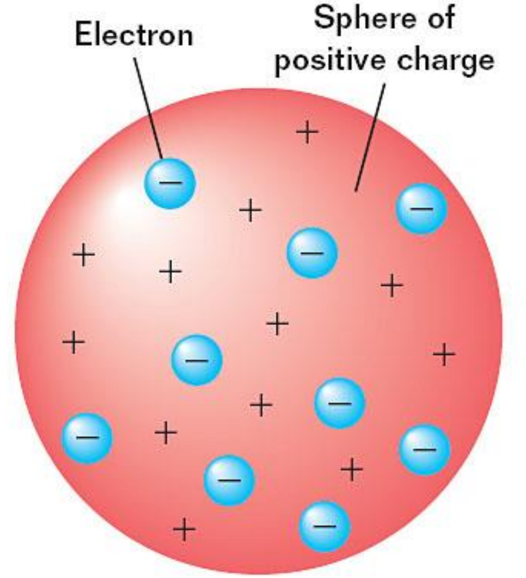
Early Models of the Atom

- The model of the atom in the days of Newton was that of a tiny, hard, indestructible sphere.
- The discovery of the electron in 1897 prompted **J. J. Thomson** (1856–1940) to suggest a new model of the atom.
- In Thomson's model, electrons are embedded in a spherical volume of positive charge like seeds in a watermelon.



The Thomson model of the atom

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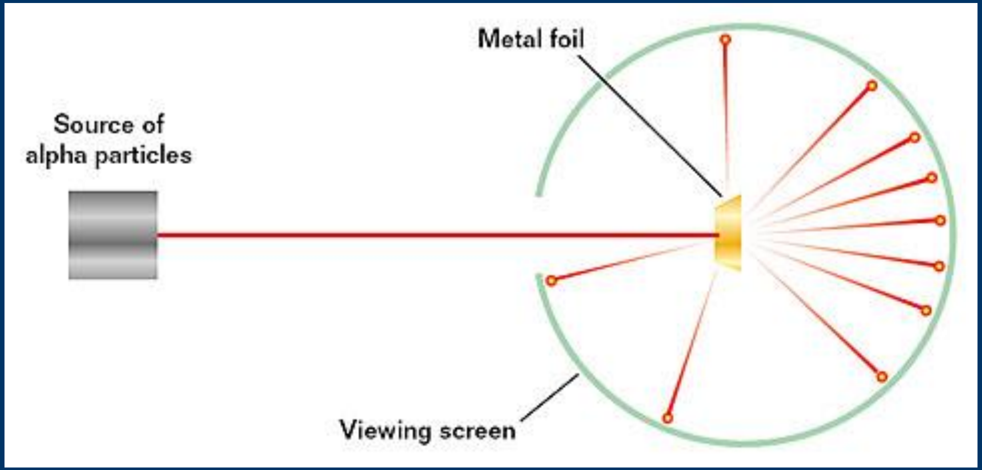
Early Models of the Atom, *continued*

- **Ernest Rutherford** (1871–1937) later proved that Thomson's model could not be correct.
- In his experiment, a beam of **positively charged alpha particles** was projected against a thin metal foil.
- Most of the alpha particles passed through the foil. Some were deflected through very large angles.



Early Models of the Atom, *continued*

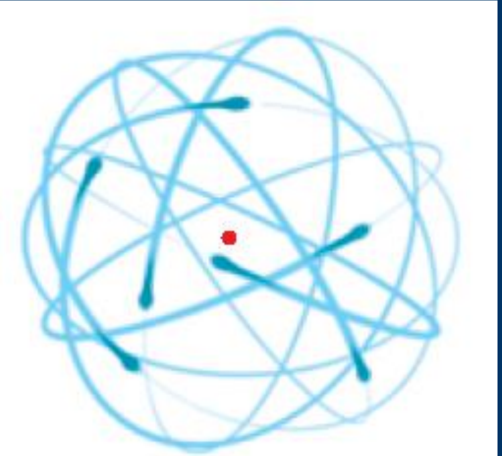
- Rutherford concluded that **all of the positive charge** in an atom and **most of the atom's mass** are found in a region that is small compared to the size of the atom.

- He called this region the the **nucleus** of the atom.

- Any **electrons** in the atom were assumed to be in the relatively large volume outside the nucleus.

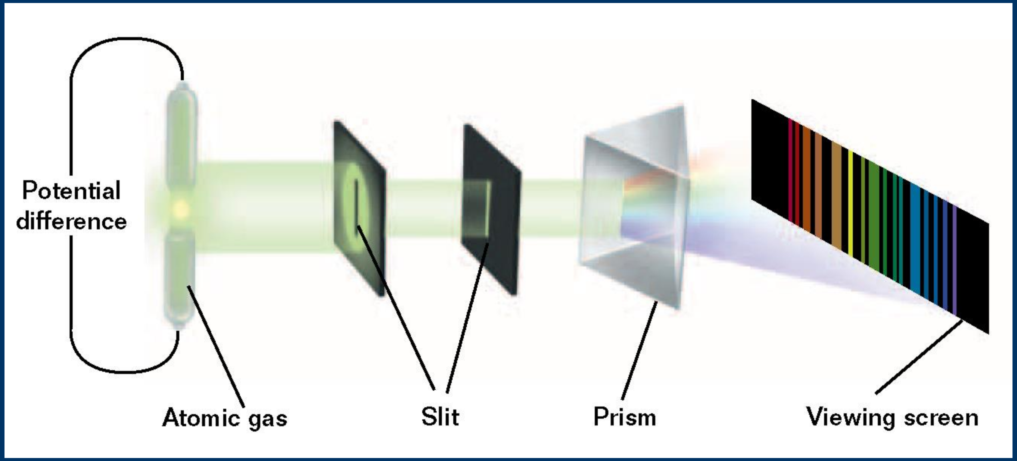
Early Models of the Atom, *continued*

- To explain why electrons were not pulled into the nucleus, Rutherford viewed the electrons as moving in orbits about the nucleus.
- However, accelerated charges should radiate electromagnetic waves, losing energy. This would lead to a rapid collapse of the atom.
- This difficulty led scientists to continue searching for a new model of the atom.


The Rutherford model

Atomic Spectra

When the light given off by an atomic gas is passed through a prism, a series of distinct bright lines is seen. Each line corresponds to a different wavelength, or color.



- A diagram or graph that indicates the wavelengths of radiant energy that a substance emits is called an **emission spectrum.**
- Every element has a **distinct emission spectrum.**

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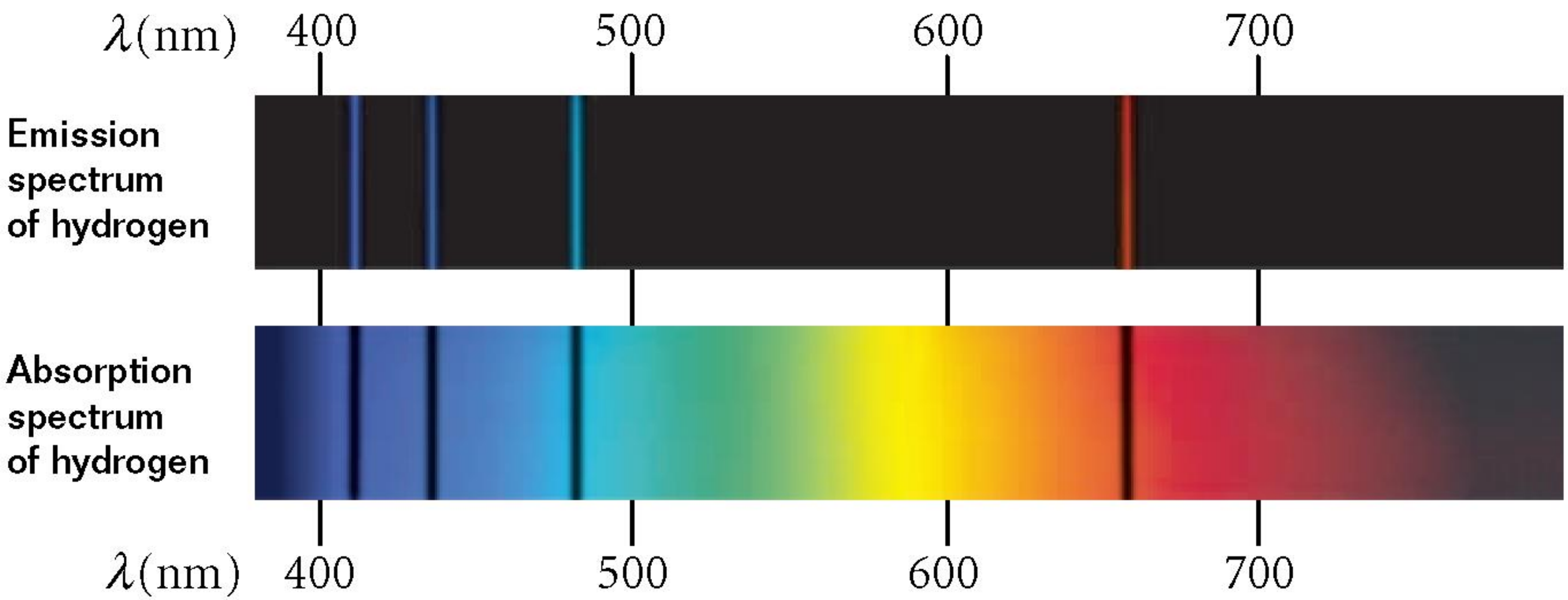
Atomic Spectra, *continued*

- An element can also **absorb** light at specific wavelengths.
- The spectral lines corresponding to this process form what is known as an **absorption spectrum.**
- An absorption spectrum can be seen by passing light containing **all wavelengths** through a vapor of the element being analyzed.
- Each line in the **absorption spectrum** of a given element coincides with a line in the **emission spectrum** of that element.

Emission and Absorption Spectra of Hydrogen

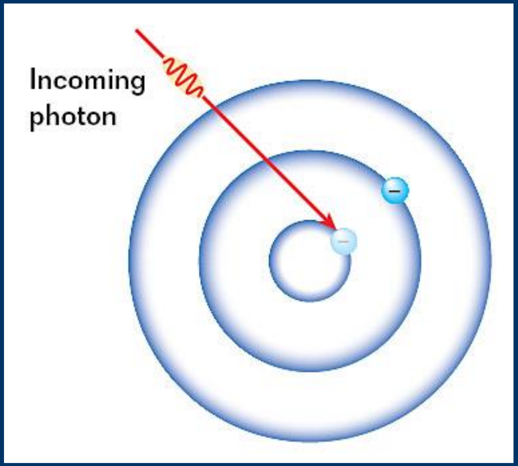


The Bohr Model of the Hydrogen Atom

- In 1913, the Danish physicist **Niels Bohr** (1885–1962) proposed a new model of the hydrogen atom that explained atomic spectra.
- In Bohr's model, **only certain orbits are allowed.** The electron is never found between these orbits; instead, it is said to "jump" instantly from one orbit to another.
- In Bohr's model, **transitions** between stable orbits with different energy levels account for the **discrete spectral lines.**

The Bohr Model, *continued*

- When light of a continuous spectrum shines on the atom, only the photons whose **energy** (***hf***) matches the **energy separation between two levels** can be absorbed by the atom.

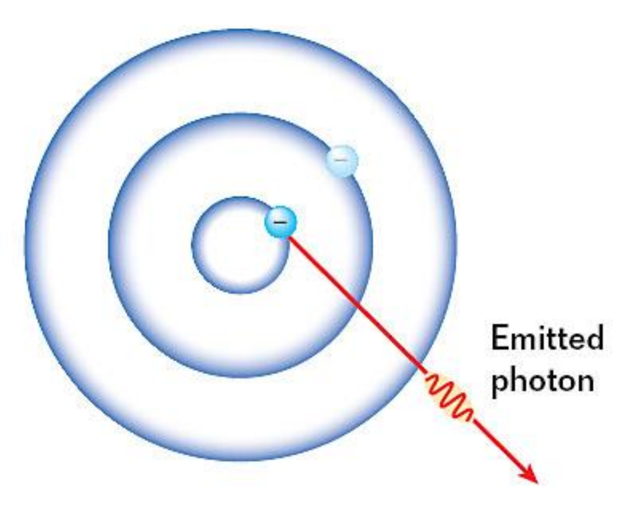


- When this occurs, an electron jumps from a **lower** energy state to a **higher** energy state, which corresponds to an orbit **farther** from the nucleus.

- This is called an **excited state**. The absorbed photons account for the dark lines in the absorption spectrum.

The Bohr Model, *continued*

- Once an electron is in an excited state, there is a certain probability that it will jump back to a **lower** energy level by **emitting a photon.**
- This process is called **spontaneous emission.**



- The emitted photons are responsible for the bright lines in the emission spectrum.
- In both cases, there is a correlation between the **"size"** of an electron's jump and the **energy** of the photon.

The Bohr Model, *continued*

- **Bohr's model** was not considered to be a complete picture of the structure of the atom.
  - Bohr assumed that electrons do not radiate energy when they are in a stable orbit, but his model offered no explanation for this.
  - Another problem with Bohr's model was that it could not explain why electrons always have certain stable orbits
- For these reasons, scientists continued to search for a new model of the atom.

The Dual Nature of Light

- As seen earlier, there is considerable evidence for the **photon** theory of light. In this theory, all electromagnetic waves consist of photons, particle-like pulses that have energy and momentum.
- On the other hand, light and other electromagnetic waves exhibit interference and diffraction effects that are considered to be **wave** behaviors.
- So, which model is correct?

The Dual Nature of Light, *continued*

- Some experiments can be better explained or only explained by the **photon** concept, whereas others require a **wave** model.
- Most physicists **accept both models** and believe that the true nature of light is not describable in terms of a single classical picture.
  - At one extreme, the electromagnetic wave description suits the overall interference pattern formed by a large number of photons.
  - At the other extreme, the particle description is more suitable for dealing with highly energetic photons of very short wavelengths.

Matter Waves

- In 1924, the French physicist **Louis de Broglie** (1892–1987) extended the wave-particle duality. De Broglie proposed that all forms of matter may have both wave properties and particle properties.
- Three years after de Broglie's proposal, C. J. Davisson and L. Germer, of the United States, discovered that electrons can be diffracted by a single crystal of nickel. This important discovery provided the **first experimental confirmation** of de Broglie's theory.

Matter Waves, *continued*

- The **wavelength** of a photon is equal to **Planck's constant** (***h***) divided by the photon's **momentum** (***p***). De Broglie speculated that this relationship might also hold for matter waves, as follows:

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

$$\text{de Broglie wavelength} = \frac{\text{Planck's constant}}{\text{momentum}}$$

- As seen by this equation, the **larger the momentum** of an object, the **smaller its wavelength.**

Matter Waves, *continued*

- In an analogy with photons, de Broglie postulated that the **frequency** of a matter wave can be found with **Planck's equation,** as illustrated below:

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

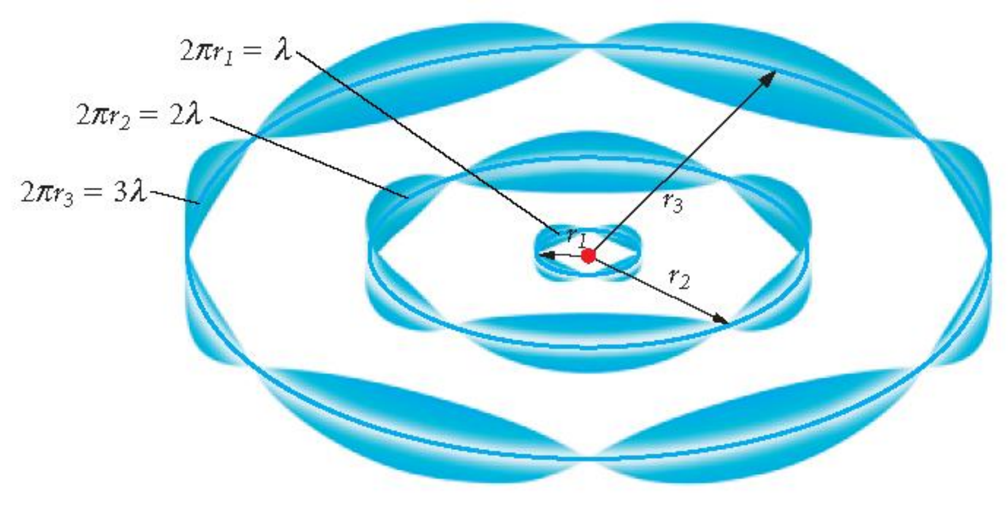
$$\text{frequency} = \frac{\text{energy}}{\text{Planck's constant}}$$

- The **dual nature of matter** suggested by de Broglie is quite apparent in the wavelength and frequency equations, both of which contain **particle concepts (*E* and *mv*)** and **wave concepts (λ and *f*)**.

Matter Waves, *continued*

- De Broglie saw a connection between his **theory of matter waves** and the **stable orbits in the Bohr model.**

- He assumed that an electron orbit would be stable only if it contained an **integral** (whole) number of electron wavelengths.



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The Uncertainty Principle

- In 1927, Werner Heisenberg argued that it is fundamentally impossible to make simultaneous measurements of a particle’s position and momentum with infinite accuracy.
- In fact, the more we learn about a particle’s momentum, the less we know of its position, and the reverse is also true.
- This principle is known as **Heisenberg’s uncertainty principle.**

The Electron Cloud, *continued*

- Quantum mechanics also predicts that the electron can be found in a **spherical region** surrounding the nucleus.
- This result is often interpreted by viewing the electron as a **cloud** surrounding the nucleus.
- Analysis of each of the energy levels of hydrogen reveals that the **most probable electron location** in each case is in agreement with each of the **radii predicted by the Bohr theory.**

The Electron Cloud

- Because the electron's location cannot be precisely determined, it is useful to discuss the **probability** of finding the electron at different locations.
- The diagram shows the probability per unit distance of finding the electron at various distances from the nucleus in the ground state of hydrogen.