

Bohr's model of hydrogen

How Bohr's model of hydrogen explains atomic emission spectra.

- **Bohr's model of hydrogen** is based on the nonclassical assumption that electrons travel in specific **shells**, or orbits, around the nucleus.
- Bohr's model calculated the following energies for an electron in the shell, n :

$$E(n) = -\frac{1}{n^2} \cdot 13.6 \text{ eV}$$

- Bohr explained the hydrogen spectrum in terms of *electrons absorbing and emitting photons* to change energy levels, where the photon energy is

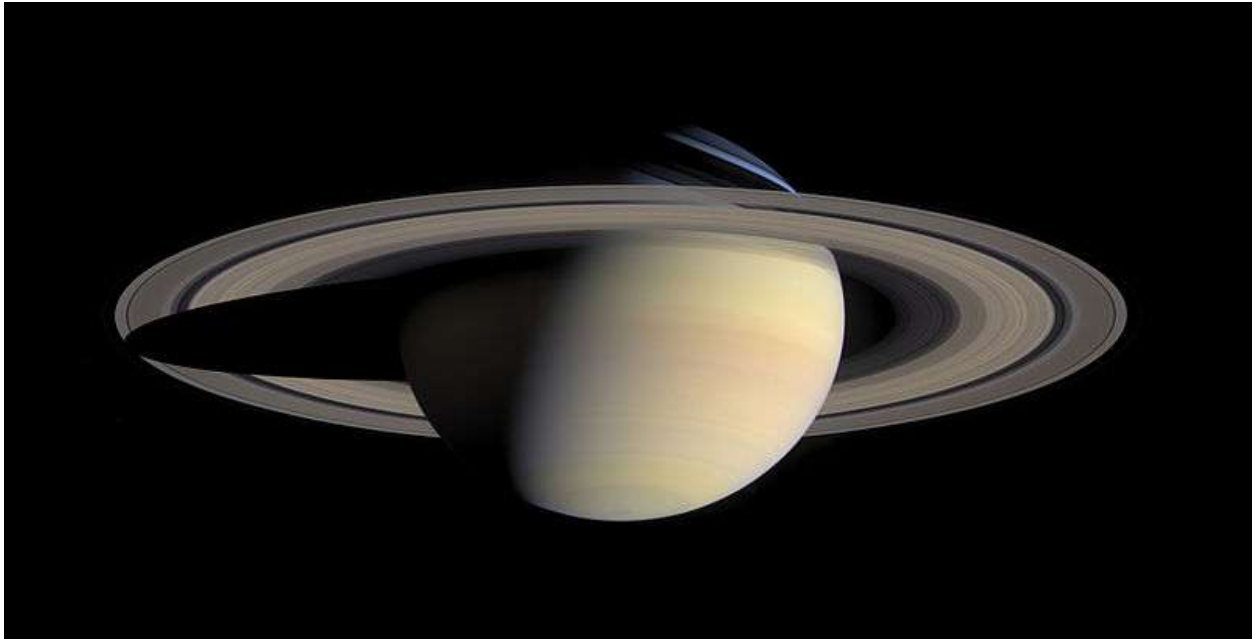
$$h\nu = \Delta E = \left(\frac{1}{n_{\text{low}}^2} - \frac{1}{n_{\text{high}}^2} \right) \cdot 13.6 \text{ eV}$$

- Bohr's model does not work for systems with more than one electron.

The planetary model of the atom

At the beginning of the 20th century, a new field of study known as quantum mechanics emerged. One of the founders of this field was Danish physicist Niels Bohr, who was interested in explaining the discrete line spectrum observed when light was emitted by different elements. Bohr was also interested in the structure of the atom, which was a topic of much debate at the time. Numerous models of the atom had been postulated based on

experimental results including the discovery of the electron by J. J. Thomson and the discovery of the nucleus by Ernest Rutherford. Bohr supported the planetary model, in which electrons revolved around a positively charged nucleus like the rings around Saturn—or alternatively, the planets around the sun.



Many scientists, including Rutherford and Bohr, thought electrons might orbit the nucleus like the rings around Saturn

However, scientists still had many unanswered questions:

- Where are the electrons, and what are they doing?
- If the electrons are orbiting the nucleus, why don't they fall into the nucleus as predicted by classical physics?
 - How is the internal structure of the atom related to the discrete emission lines produced by excited elements?

Bohr addressed these questions using a seemingly simple assumption: what if some aspects of atomic structure, such as electron orbits and energies, could only take on certain values?

Quantization and photons

By the early 1900s, scientists were aware that some phenomena occurred in a discrete, as opposed to continuous, manner. Physicists Max Planck and Albert Einstein had recently theorized that electromagnetic radiation not only behaves like a wave, but also sometimes like particles called *photons*. Planck studied the electromagnetic radiation emitted by heated objects, and he proposed that the emitted electromagnetic radiation was "quantized" since the energy of light could only have values given by the following equation:

$$E_{\text{photon}} = nh\nu, \text{ where } n \text{ is a positive integer, } h \text{ is Planck's constant—} 6.626 \times 10^{-34} \text{ J} \cdot \text{s} \text{—and } \nu \text{ is the frequency of the light, which has units of } \frac{1}{\text{s}}.$$

As a consequence, the emitted electromagnetic radiation must have energies that are multiples of $h\nu$. Einstein used Planck's results to explain why a minimum frequency of light was required to eject electrons from a metal surface in the photoelectric effect.

When something is **quantized**, it means that only specific values are allowed, such as when playing a piano. Since each key of a piano is tuned to a specific note, only a certain set of notes—which correspond to frequencies of sound waves—can be produced. As long as your piano is properly tuned, you can play an F or F sharp, but you can't play the note that is halfway between an F and F sharp.

Atomic line spectra

Atomic line spectra are another example of quantization. When an element or ion is heated by a flame or excited by electric current, the excited atoms emit light of a characteristic color. The emitted light can be refracted by a prism, producing spectra with a distinctive striped appearance due to the emission of certain wavelengths of light.

For the relatively simple case of the hydrogen atom, the wavelengths of some emission lines could even be fitted to mathematical equations. The equations did not explain why the hydrogen atom emitted those particular wavelengths of light, however. Prior to Bohr's model of the hydrogen atom, scientists were unclear of the reason behind the quantization of atomic emission spectra.

Bohr's model of the hydrogen atom: quantization of electronic structure

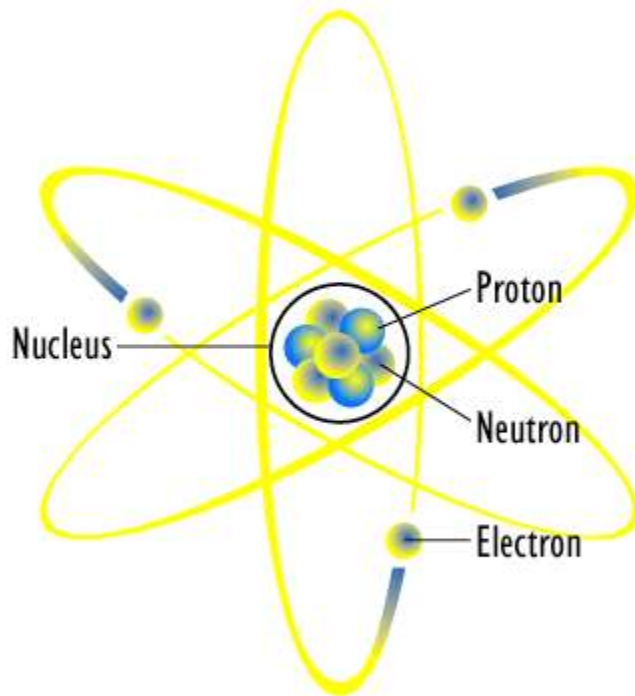
Bohr's model of the hydrogen atom started from the planetary model, but he added one assumption regarding the electrons. What if the electronic structure of the atom was quantized? Bohr suggested that perhaps the electrons could only orbit the nucleus in specific orbits or *shells* with a fixed radius. Only shells with a radius given by the equation below would be allowed, and the electron could not exist in between these shells.

Mathematically, we could write the allowed values of the atomic radius as

$r(n) = n^2 \cdot r(1)$, where n is a positive integer, and $r(1)$ is the **Bohr radius**, the smallest allowed radius for hydrogen.

He found that $r(1)$ has the value

$$\text{Bohr radius} = r(1) = 0.529 \times 10^{-10} \text{ m}$$

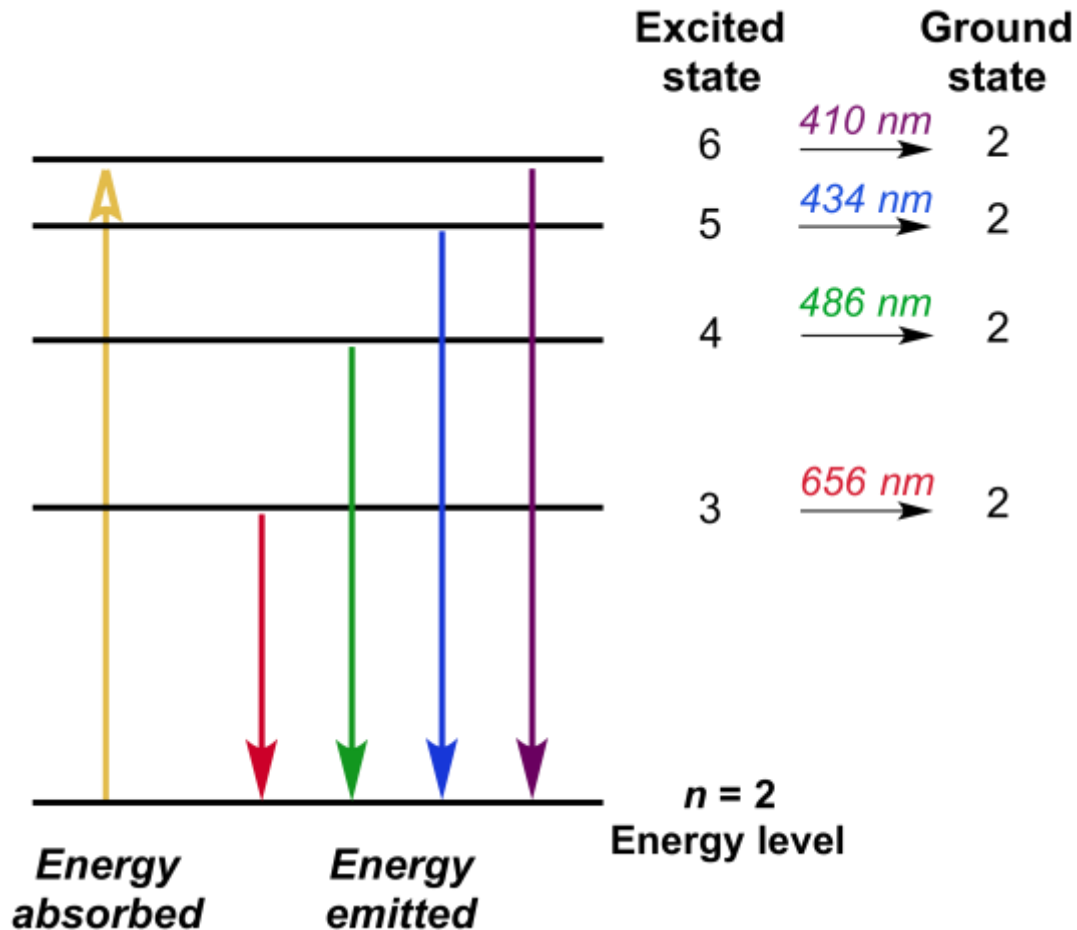


An atom of lithium shown using the planetary model. The electrons are in circular orbits around the nucleus.

By keeping the electrons in circular, quantized orbits around the positively-charged nucleus, Bohr was able to calculate the energy of an electron in the n th energy level of hydrogen: $E(n) = -\frac{1}{n^2} \cdot 13.6 \text{ eV}$, where the lowest possible energy or **ground state energy** of a hydrogen electron— $E(1)$ —is -13.6 eV .

Note that the energy is always going to be a negative number, and the ground state, $n = 1$, has the most negative value. This is because the energy of an electron in orbit is relative to the energy of an electron that has been completely separated from its nucleus, $n = \infty$, which is defined to have an energy of 0 eV . Since an electron in orbit around the nucleus is more stable than an electron that is infinitely far away from its nucleus, the energy of an electron in orbit is always negative.

Absorption and emission



The Balmer series—the spectral lines in the visible region of hydrogen's emission spectrum—corresponds to electrons relaxing from $n=3-6$ energy levels to the $n=2$ energy level.

Bohr could now precisely describe the processes of absorption and emission in terms of electronic structure. According to Bohr's model, an electron would absorb energy in the form of photons to get excited to a higher energy level *as long as the photon's energy was equal to the energy difference between the initial and final energy levels*. After jumping to the higher energy level—also called the **excited state**—the excited electron would be in a less stable position, so it would quickly emit a photon to relax back to a lower, more stable energy level.

The energy levels and transitions between them can be illustrated using an **energy level diagram**, such as the example above showing electrons relaxing back to the $n = 2$ level of hydrogen. The energy of the emitted photon is equal to the difference in energy between the two energy levels for a particular transition. The energy difference between energy levels n_{high} and n_{low} can be calculated using the equation for $E(n)$ from the previous section:

$$\begin{aligned}\Delta E &= E(n_{high}) - E(n_{low}) \\ &= \left(-\frac{1}{n_{high}^2} \cdot 13.6 \text{ eV} \right) - \left(-\frac{1}{n_{low}^2} \cdot 13.6 \text{ eV} \right) \\ &= \left(\frac{1}{n_{low}^2} - \frac{1}{n_{high}^2} \right) \cdot 13.6 \text{ eV}\end{aligned}$$

Since we also know the relationship between the energy of a photon and its frequency from Planck's equation, we can solve for the frequency of the emitted photon:

$$h\nu = \Delta E = \left(\frac{1}{n_{low}^2} - \frac{1}{n_{high}^2} \right) \cdot 13.6 \text{ eV} \quad \text{Set photon energy equal to } \Delta E$$

$$\nu = \left(\frac{1}{n_{low}^2} - \frac{1}{n_{high}^2} \right) \cdot \frac{13.6 \text{ eV}}{h} \quad \text{Solve for frequency}$$

We can also find the equation for the wavelength of the emitted electromagnetic radiation using the relationship between the speed of light c , frequency ν , and wavelength λ :

$$c = \lambda\nu$$

Rearrange to :

$$\frac{c}{\lambda} = \nu = \left(\frac{1}{n_{low}^2} - \frac{1}{n_{high}^2} \right) \cdot \frac{13.6 \text{ eV}}{h}$$

Divide both sides

$$\frac{1}{\lambda} = \left(\frac{1}{n_{low}^2} - \frac{1}{n_{high}^2} \right) \cdot \frac{13.6 \text{ eV}}{hc}$$

Thus, we can see that the frequency—and wavelength—of the emitted photon depends on the energies of the initial and final shells of an electron in hydrogen.

What have we learned since Bohr proposed his model of hydrogen?

The Bohr model worked beautifully for explaining the hydrogen atom and other single electron systems such as He^+ . Unfortunately, it did not do as well when applied to the spectra of more complex atoms. Furthermore, the Bohr model had no way of explaining why some lines are more intense than others or why some spectral lines split into multiple lines in the presence of a magnetic field—the Zeeman effect.

In the following decades, work by scientists such as Erwin Schrödinger showed that electrons can be thought of as behaving like waves *and* behaving

as particles. This means that it is not possible to know both a given electron's position in space and its velocity at the same time, a concept that is more precisely stated in *Heisenberg's uncertainty principle*. The uncertainty principle contradicts Bohr's idea of electrons existing in specific orbits with a known velocity and radius. Instead, we can only calculate *probabilities* of finding electrons in a particular region of space around the nucleus.

The modern quantum mechanical model may sound like a huge leap from the Bohr model, but the key idea is the same: classical physics is not sufficient to explain all phenomena on an atomic level. Bohr was the first to recognize this by incorporating the idea of quantization into the electronic structure of the hydrogen atom, and he was able to thereby explain the emission spectra of hydrogen as well as other one-electron systems.