

Lesson 40: Bohr's Model

While Rutherford was working in his lab at the University of Manchester trying to figure out atoms, he had a person named [Niels Bohr](#) working for him.

- Previously Bohr had worked for J.J. Thomson.
- Bohr showed that he was a better than average physicists by suggesting multiple improvements to the experiments being done in Rutherford's lab.
- Eventually Bohr was even able to suggest an answer to the problem with Rutherford's Planetary model.



Illustration 1: Niels Bohr.

Bohr was able to take two ideas and put them together to come up with his model.

1. Rutherford's Planetary Model

Bohr knew that he couldn't completely trash Rutherford's Planetary model.

- The experiments Rutherford had done proved that there must be a positive nucleus surrounded by negative electrons.
- The change Bohr came up with concerned the electrons' positions.
 - Rather than allowing the electrons to be just anywhere, he suggested that they could only exist at specific distances from the nucleus.

2. Quantum Mechanics' Discrete Energy

Specific positions for the electrons agreed nicely with the ideas of discrete amounts of energy that we learned about in quantum mechanics.

- As we learned in the last lesson, if the electrons were anywhere they would eventually spiral in and crash.
 - In fact, as they spiral in they should go faster and faster and release their energy in ever increasing amounts.
 - Since $E = hf$, this would mean we should see a bunch of increasingly higher frequencies of EMR being given off by atoms all the time. But we don't.
- Instead, experiments showed that atoms only ever emit energy at specific frequencies.
 - This must mean that specific, discrete amounts of energy are being emitted.

Spectroscopy

There was strong evidence of energy being released as quanta when physicists studied **spectroscopy**. This was one of Bohr's inspirations to come up with his model, as well as some of its best support.

- Spectroscopy involves looking at light from various sources through a diffraction grating and analyzing the colors that are seen.

Emission Spectra

In experiments it was discovered that if a high voltage (or heat) is applied to a gas sample of an element in a tube at low pressure, it will give off light.

- When a diffraction grating is used to examine the light, it does not emit a **continuous spectrum** of colors. Instead, each element releases an **emission line spectrum** made up of individual lines of color at specific frequencies.
 - Every element gives off its own unique emission line spectrum. It acts like a fingerprint of that element.

A **continuous spectrum** is when pure white light is broken up to show all the colors of the rainbow without any missing.

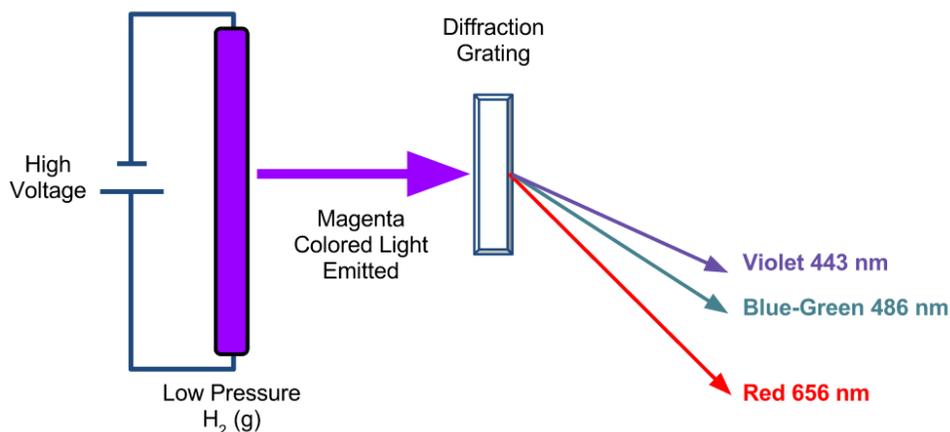


Illustration 2: Set up for producing an emission spectra.

Absorption Spectra

The opposite is also true.

- If we shine a continuous spectrum of white light through a cool gas at low pressure, it will absorb the light at the same frequencies as it emits.
 - These missing lines from the continuous spectra are called the **absorption line spectrum**.

As early as 1814 Josef von Fraunhofer had noticed that if you allowed sunlight to go through a prism, the resulting spectrum of visible light had some black lines in it. This absorption line spectrum lets us know what the Sun is made from.

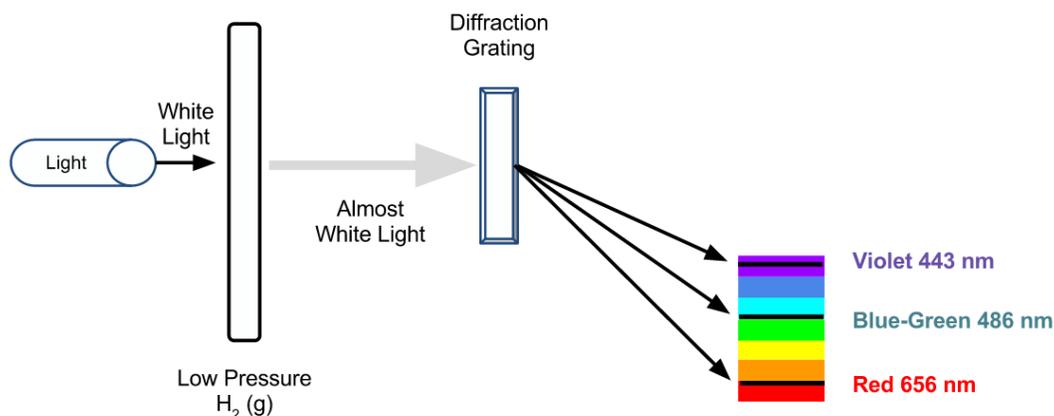


Illustration 3: Set up to produce an absorption spectra.

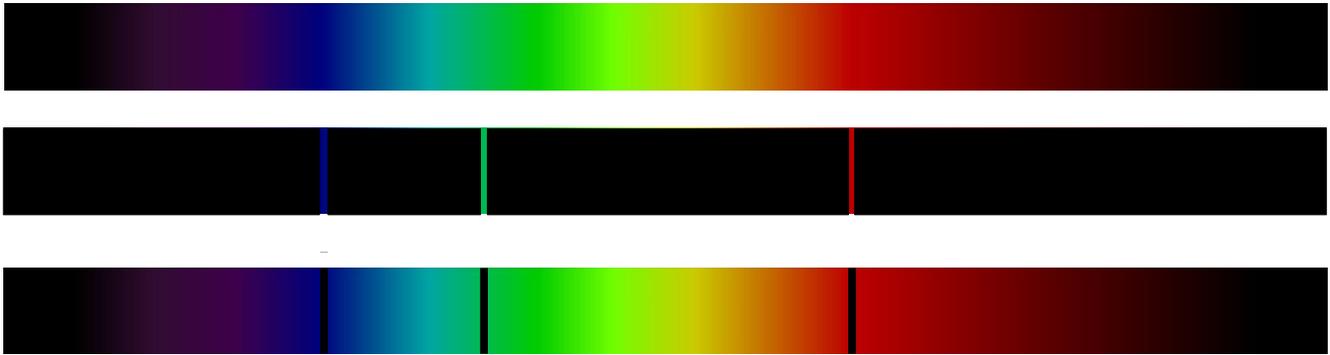


Illustration 4: A continuous spectrum shows all the colors of the rainbow (like you get from heating a solid), an emission line spectrum shows just a few slivers of color, and an absorption line spectrum shows a few black slivers missing.

Balmer's Formula

A Swiss high school teacher [Johann Jakob Balmer](#) came up with a formula that showed that the spectra being seen in these experiments followed a clear pattern.

- Johannes Robert Rydberg improved this formula and came up with a standard version for hydrogen that let you calculate the wavelengths of light emitted.
 - Hydrogen is important in that it is the simplest element, and scientists were focusing on explaining it first.

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

λ = wavelength of light emitted (m)
 R_H = Rydberg's Constant for Hydrogen
 n_f = final energy level
 n_i = initial energy level

When Bohr saw this formula he was inspired, because everything fit together so well.

- The idea of electrons existing at specific positions shows up in the formula as “n” values.
 - Because an electron would have a specific amount of energy at these particular positions, we will call them **energy levels**.
- Bohr said that electrons in these energy levels are stable and do not emit EMR as they orbit.
 - Since these electrons are **not** losing energy, we can also refer to these positions as **stationary states**. This does not mean that the electrons are stationary, just that they are in a stable position.
- Bohr theorized that energy in the form of EMR is...
 - ...**emitted** when an electron **falls down** from one level to a lower level.
 - ...**absorbed** to make an electron **jumps up** from one level to a higher level.

Example 1: Determine the wavelength of light emitted when an electron in hydrogen jumps from the second excited state (the third energy level) to the ground state.

The ground state is a name used to refer to the closest the electron can get to the nucleus, $n = 1$. The electron can't fall any further, so it's like it hit the ground.

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

$$\frac{1}{\lambda} = 1.10 \times 10^7 \left(\frac{1}{1^2} - \frac{1}{3^2} \right)$$

$$\lambda = 1.02272 \times 10^{-7} \text{ m} = 1.02 \times 10^{-7} \text{ m}$$

There are names given to some specific jumps in hydrogen based on the names of the people that first studied them.

- Lyman Series**
 Measured by Theodore Lyman.
 Always has a final energy level of 1.
 This is such a large fall that it releases lots of energy and UV light is emitted.
- Balmer Series**
 Measured by Balmer himself.
 Always has a final energy level of 2.
 Since this is not as great a fall, not as much energy is released. This is why visible light is seen.
- Paschen Series**
 Louis Paschen measured this one.
 It always has a final level of 3, so it releases very little energy.
 It is always infrared radiation emitted.

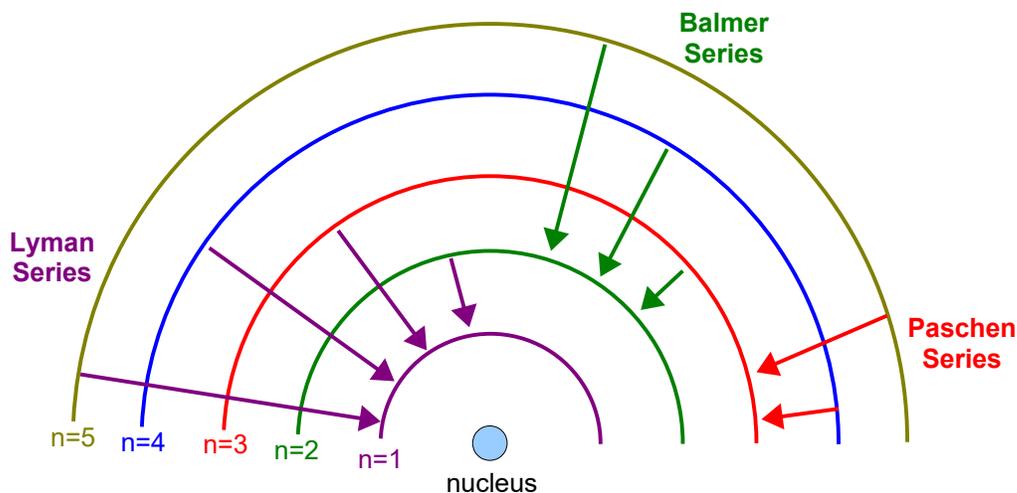


Illustration 3: The Lyman, Balmer, and Paschen series all have different stopping points for the electron. This diagram is not drawn to scale.

Example 2: Determine the (a) longest and (b) shortest wavelengths possible in the Paschen series. The Paschen series always has $n_f = 3$.

a) The longest wavelength is the smallest frequency possible. Since $E = hf$, a small frequency means the released EMR will have a small amount of energy. If only a small amount of energy is released, the electron must have only fallen the smallest amount possible. This means the electron must have started just above the third energy level at the fourth energy level.

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

$$\frac{1}{\lambda} = 1.10e7 \left(\frac{1}{3^2} - \frac{1}{4^2} \right)$$

$$\lambda = 1.87012987e-6 = 1.87e-6 \text{ m}$$

b) The shortest wavelength is the biggest frequency possible. Since $E = hf$, a big frequency means the released EMR will have a huge amount of energy. If a huge amount of energy is released, the electron must have only fallen a long ways. What's the highest up in can start? Well, technically, $n = \text{infinity!}$ So we will set $n_i = \infty$.

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

$$\frac{1}{\lambda} = 1.10e7 \left(\frac{1}{3^2} - \frac{1}{\infty^2} \right)$$

$$\frac{1}{\lambda} = 1.10e7 \left(\frac{1}{3^2} - 0 \right)$$

$$\lambda = 8.18e-7 = 8.18e-7 \text{ m}$$

Energy Level Sizes and Energies

Using other formulas Bohr calculated two special numbers for the lowest energy level ($n=1$) of hydrogen

- He figured out the **radius** of the orbit of electrons in the lowest energy level for hydrogen as being $5.29e-11 \text{ m}$
- He also figured out the **energy** in the ground state as $-2.18e-18 \text{ J}$ (-13.6 eV).
 - Bohr made it negative because you must add energy to jump it up. 13.6 eV is the amount of energy you would have to add to an electron starting in $n=1$ to be able to get it to be ionized out of the hydrogen.

Using these numbers for the ground state you can use two formulas Bohr came up with to figure out...

- the radius of the n^{th} energy level.

$$r_n = n^2 r_1$$

r_n = radius of n^{th} energy level (m)
 n = energy level
 r_1 = radius of first energy level (m)

- the energy of the n^{th} energy level.

$$E_n = \frac{1}{n^2} E_1$$

E_n = energy of n^{th} energy level (J or eV)

n = energy level

E_1 = energy of first energy level (-2.18e-18 J or -13.6eV)

Bohr thought these formulas were great!

- Not only are they useful, they also agree with everything else in the model.
 - You can even use these formulas to prove that Balmer's formula is correct.
- Bohr also thought it was good that they were such easy formulas, since simple explanations are usually correct in science.
- The energies are negative, since Bohr pictured the electrons as though they were at the bottom of a hole. By adding that energy we can lift the electron out (ionizing the atom).

Example 3: Determine the (a) radius and (b) energy of an electron in the fourth energy level of hydrogen.

$$r_n = n^2 r_1$$

a) $r_4 = 4^2 (5.29\text{e-}11)$

$$r_4 = 8.464\text{e-}10 = 8.46\text{e-}10 \text{ m}$$

$$E_n = \frac{1}{n^2} E_1$$

b) $E_4 = \frac{1}{4^2} (-2.18\text{e-}18)$

$$E_4 = -1.3625\text{e-}19 = -1.36\text{e-}19 \text{ J}$$

Note, this answer could also be given in eV.

Unfortunately *all* these formulas only work for the simplest atom, hydrogen.

- If you try to apply the formulas to any other element, you get the answers that do not agree with actual experiments.
- Bohr understood that there were problems with his model, but he hoped that there might be a way to improve it.
- It is still valid to use it to make predictions.

Did YOU know?

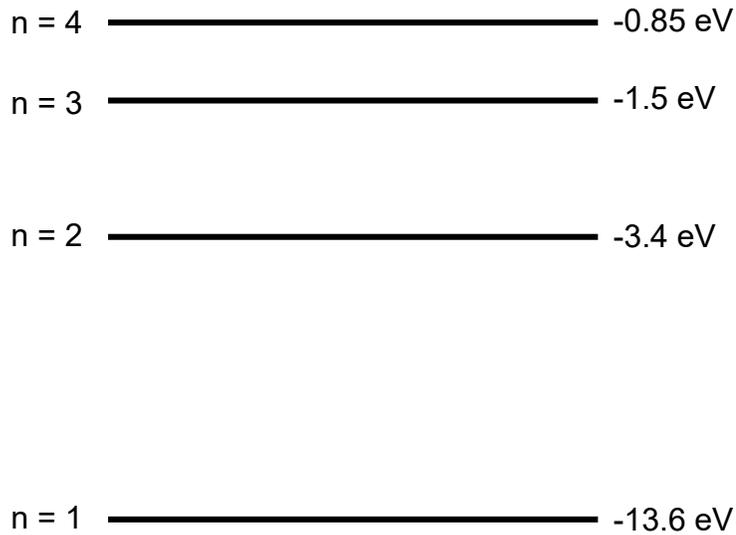
Bohr once said "Your theory is crazy, but it's not crazy enough to be true."

Jumping Energy Levels

Jumping through energy levels means energy is either absorbed or released.

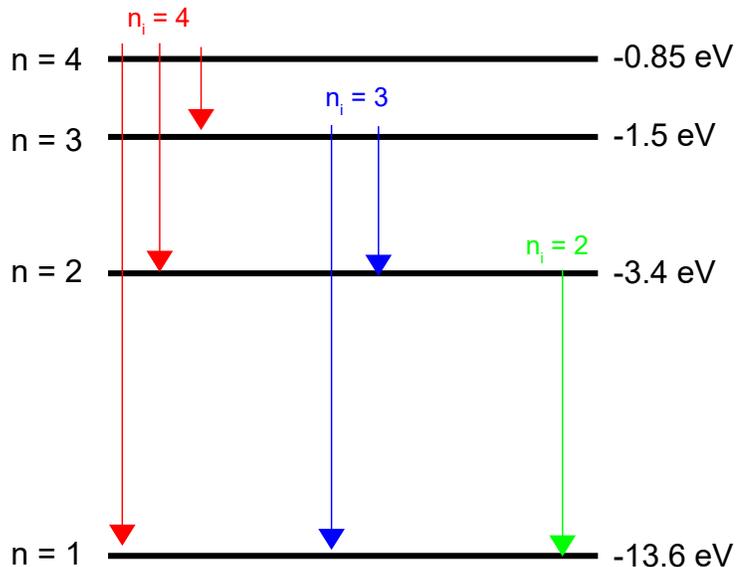
- Lots of different, specific amounts of energy can be involved, since many different combinations of jumps can occur.
- The energy involved is always equal to the difference between the two specific energy levels.

Example 4: The following diagram (not drawn to scale) shows four energy levels of hydrogen.



- a) **Determine** the total number of possible energies emitted as electrons fall through these levels.
b) **Determine** the wavelength of EMR emitted as an electron drops from n_3 to n_1 .

a) Sketch out all the possible “falls” the electron can go through. It doesn't have to fall all the way down, which is why we can have many different possible energies.



Count 'em up! This gives us a total of **6 possible energies** that will be released.

- b) Each of the falls we just looked at involves the electron releasing a specific amount of energy; each of these energies is released as a specific frequency/wavelength of EMR. We can figure out the energy difference between n_3 to n_1 , and then calculate the wavelength.

$$\begin{aligned}\Delta E &= E_f - E_i \\ \Delta E &= E_1 - E_3 \\ \Delta E &= (-13.6) - (-1.5) \\ \Delta E &= -12.1 \text{ eV} \\ E &= \frac{hc}{\lambda} \rightarrow \lambda = \frac{hc}{E} \\ \lambda &= \frac{4.14 \times 10^{-15} (3.00 \times 10^8)}{12.1} \\ \lambda &= 1.02645 \times 10^{-7} = 1.03 \times 10^{-7} \text{ m}\end{aligned}$$

We use Planck's constant in eV·s since we have an energy in eV. Also, we use the energy as an absolute value, since we know this is energy being released as EMR.

Homework

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p.778 #2
p.780 #3, 5, 11