

8.4 The Bohr Atom pg 290

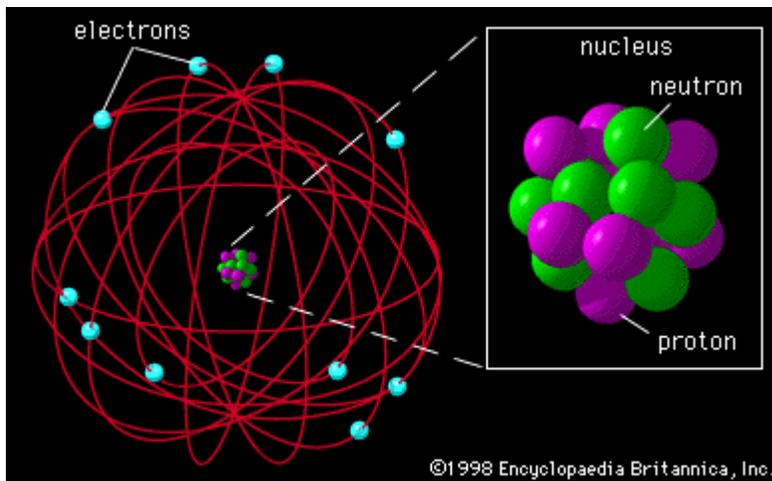
Key equations:

$$E_n = \frac{-Z^2 hc R_h}{n^2}$$
$$\Delta E = -hc R_h Z^2 \left(\frac{1}{n_f^2} - \frac{1}{n_i^2} \right)$$
$$r_n = \frac{n^2 a_0}{Z}$$

This should be fun. 8.4 is not that long.

Ok, before we pull the Bohr atom out of no where, lets start with the Rutherford atom. Rutherford's model of an atom is one of the earliest models of the atom. It basically said that there are electrons that circle around a nucleus like planets that circle around the sun. Well, in an electron is a charged object (it has a negative charge) and it is accelerating (in physics, acceleration occurs when direction changes or speed changes. It is a difficult concept. It is hard to imagine that something that runs in a circle would accelerate but it does accelerate!)

Ok, so you have a charged object that is continuously accelerating. (since it is traveling continuously in a circle around the nucleus) What happens is that as electrons accelerate around the nucleus, they should continuously emit energy. Since an electron contains a finite amount of energy, it will eventually lose all energy and collapse to the nucleus. This would cause the atom to collapse as well. Well, this doesn't happen or nothing would exist at all! Rutherford's atom is obviously wrong. (Oh, I don't know why electrons radiate energy when they accelerate, I would assume that it is a property of the charged particles when they accelerate. This will require some high level physics so I can't explain it)



Rutherford Atom

As I mentioned in 8.1, part of the reason we study quantum mechanics is to gain an understanding of the structure of the atom. So, let's talk about Mr. Bohr.

Mr. Bohr, very smart guy and very creative. He doesn't like to follow the laws of physics. He assumed that nature made special exceptions for an atom. Let's look at what he suggests.

- 1) Electron travels in circular orbits around the nucleus (Hmm... ok.. what's an orbit? Orbit is just some "track" that the electron moves on)
- 2) Electrons, because they are they are running in a circle, have what we called an angular momentum. (Angular momentum is not something you should worry about. This is a rather difficult concept that is only taught in Year 2 physics. Just know what angular momentum leads to. For now, angular momentum describes rotational motion about an axis.) Bohr assumed that the angular momentum is quantized and equals to $\frac{nh}{2\pi}$. This quantized angular momentum means that electrons have only a fixed set of allowed orbits, which he called *stationary states*. As long as an electron remains in a given orbit, its energy is constant and no energy is emitted. (Here, he is violating the laws of classical physics but it is the assumption he has to make to ensure that his model works) Energy can be only absorbed/emitted when electrons move to other orbits
- 3) Electrons can pass only from one allowed orbit to another (Nothing in between) by absorbing or emitting a photon whose energy equals the difference in energy between the two orbits.

$$E_{\text{photon}} = E_{\text{orbitA}} - E_{\text{orbitB}} = h\nu$$

What ideas we discussed before is incorporated in Bohr's atom?

Well, he took the idea of quantized energy suggested by Planck. This model explains the atomic spectra we discussed in 8.2. Since electrons can only absorb specific energies to jump around, they will only emit specific wavelengths of EMR that they absorbed. The idea of photons is also incorporated in Bohr's atom but it does not play a significant role. (Well, from here on, we will assume that light/EMR behaves as particles that we call photon. This is valid until we get to 8.6)

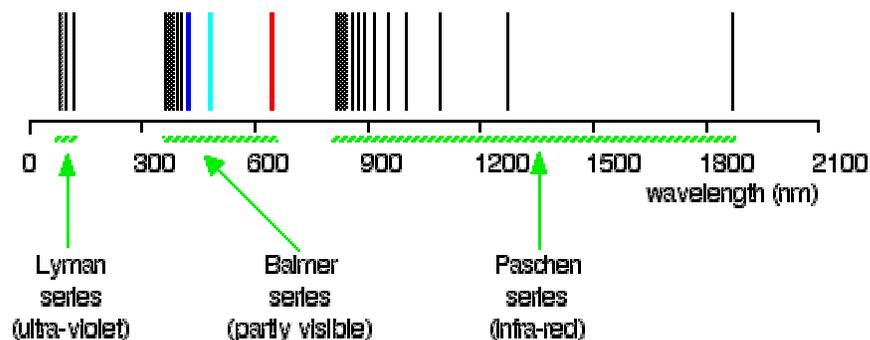
In Bohr's model, the quantum number n (1, 2, 3, 4...) is associated with the radius of an electron orbit, which is directly related to the electron's energy. (This is seen $r_n = \frac{n^2 a_0}{z}$ and in $E_n = -hcR_h(\frac{1}{n^2})$) The lower the n number, the lower the energy and smaller the radius.

Lets talk about hydrogen. Hydrogen, as you know, have only one electron. Therefore, when that single electron is in $n = 1$, the hydrogen atom is in the ground state. When that electron receives a photon that contains the energy that matches the energy difference between say $n = 2$ orbit and $n = 1$ orbit, it will move up to the $n = 2$ orbit. (Relates to Bohr's 3rd assumption, the idea that electrons can only jump between orbits if they receives a specific quantized amount of energy) When that electron is in $n \geq 2$ orbits, that hydrogen atom is said to be in an excited state. This ground state and excited state theory applies to all hydrogen like species (Atoms that only have 1 electron, i.e. He^{+1} , Li^{+2})

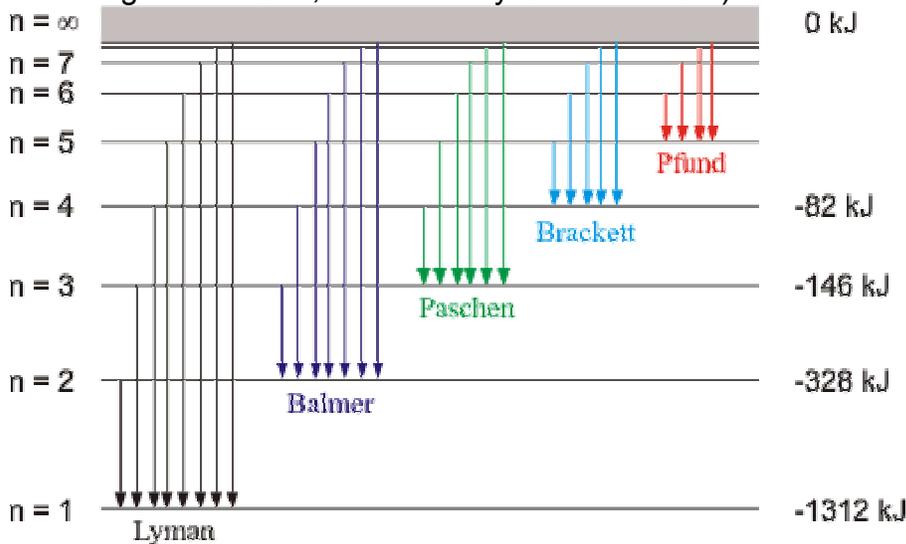
There are three spectra series you should know. What is a spectra series? You know the atomic spectra for hydrogen right? It has 4 spectra lines with specific wavelengths. (Spectra lines are the single colored lines you see in section 8.2) These 4 spectra lines make up a single spectra series. This specific series is called the Balmer series because it is in the visible spectrum of the EMR spectra. Spectra lines in the Balmer series all have a common characteristic: they are all created by electrons dropping from higher orbits into the $n = 2$ orbit. ($n_f = 2$)

The other two spectra series are the Lyman and Paschen series. Lyman series is in the UV region of the EMR spectra. Spectra lines in the Lyman series are created by electrons dropping from higher orbits into the $n = 1$ orbit. (Think about it, it makes sense that Lyman series is in the UV region since the energy release by electrons jumping from higher orbits to $n = 1$ is much higher than those that jump to $n = 2$) Paschen series is in IR spectrum. Spectra lines in the Paschen series are created when electrons drop from higher orbits to $n = 3$ orbit. (This also makes sense since electrons dropping to $n = 3$ are going to emit less energy than those that drop to $n = 2$ or $n = 1$. Thus, the EMR emitted are all the IR region, which has lower energy than UV or the visible region of the EMR spectrum.)

Here are all the spectra series placed on the EMR spectrum.



Here are all the jumps I mentioned. As you can see, the gaps between energy levels are large when the n numbers are small. As n increases, the gap decreases. This means that the jump from $n = 2$ to $n = 1$ would release much more energy than the jump from $n = 4$ to $n = 3$. (from 2 to 1, you would release $1312 - 328$ kJ which is close to 1000 kJ of energy. From 4 to 3, you are only releasing $146 - 82$ kJ, which is only around 60 kJ.)

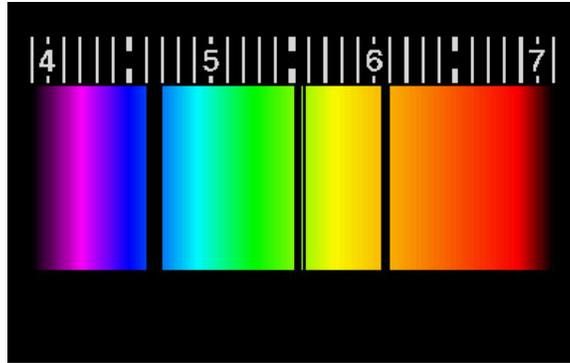


A common assumption that is often made is that these spectra series only apply when electrons emit energy. This is a wrong assumption. These series also exist when electrons absorb energy. Thus, this is why the spectra series are often called emission/absorption spectra series.

Please note that Balmer/Lyman/Paschen series only applies to hydrogen atoms. It does not apply to hydrogen like species!

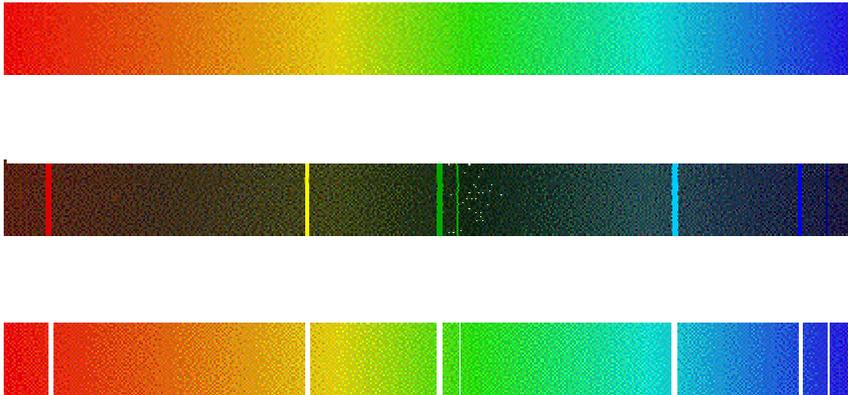
I noticed you had some confusion over ground state and Balmer/Lyman/Paschen series. I hope what I wrote so far cleared this confusion.

It is useful at this point to take a look at the absorption process. In 8.1, we have discussed in detail about the emission spectra. There is, however, another way for spectra to be created. Suppose instead of passing electricity through a glass tube containing gas, we pass white light (containing all the visible wavelengths) through a glass tube containing gaseous atoms. Then we use a prism to disperse the light that has passed through the glass tube. Instead of seeing the complete continuous spectrum, we would observe dark lines on the spectrum.



Why does this happen? When white light passes through the glass tube, some of it interacts with the atoms. If the frequency of one of the lights matches the frequency required by the gaseous atom to bump electrons up, that frequency of light will be absorbed. Thus, when you pass the remaining light through a prism (which disperse white light into the individual wavelengths that make up the white light), you will see that frequency of light “missing” from the color spectra.

The interesting thing about this is that emission spectra and absorption spectra matches! Lets take a look.



On the very top, you have the continuous spectrum. In the middle, you have the emission spectra and on the bottom you have the absorption spectra. Notice how the lines on the emission and absorption matches? They all occur at the same wavelengths! This proves that atoms really do have orbits that are quantized. It also reinforces the fact that Balmer/Lyman/Paschen series do exist for the absorption spectra.

So far we have looked at Bohr’s atom qualitatively, let’s look at it quantitatively. This means that the next couple pages are going to be heavy with equations and such but there is no need to panic. If you are solid with the theory, the equations won’t scare you.

Lets start by discussing the radius of orbits in Bohr's model of the atom. There is a simple equation you have to know, this equation is $r_n = n^2 a_0$, where n corresponds to the energy levels within the atom. This equation is only valid for hydrogen atoms. Thus, for hydrogen like species we modify it a bit so it becomes..

$$r_n = \frac{n^2 a_0}{Z}, \text{ where } Z \text{ equals to the atomic number of the hydrogen like element.}$$

Notice that it is Z and NOT Z^2 !

The energy of each orbit within the atom can also be calculated. I have not expressed this specifically before so I will do it now. Electrons that exist in a certain orbit must have energies that match exactly the energy specified by the orbit. If an electron does not have that specific energy, it cannot exist there. It is possible to calculate the energy that an electron must possess in order to exist in a specific orbit.

The formula for this is:

$$E_n = -hcR_h \left(\frac{1}{n^2} \right) \text{ (For hydrogen atoms only)}$$

$$E_n = -Z^2 hcR_h \left(\frac{1}{n^2} \right) \text{ (For hydrogen like species)}$$

Note that it is Z^2 in this equation.

The energy you calculated will be negative. This is a rather difficult concept to master. Think of the energy in each orbit as negative temperatures. At $n = 1$, it is -100 degrees. At $n = 2$, it is -50 degrees. At $n = 3$ it is -30 degrees etc.. As you go up, the temperature increase even though they are still negative. The same thing applies with energy of orbits. At the very top, when $n = \text{infinity}$, the energy is 0. When $n = \text{infinity}$, it means that the electron has been ionized from the atom. (Completely removed from the nucleus) Thus, it no longer has any energy. (The reason for this has to do with potential energy, which is a concept in physics. We will not discuss it here but if you want to know, just ask me)

Armed with the knowledge of calculating energies of each orbit, we can calculate the energy differences between orbits. This will allow us to determine how much energy an electron needs to be able to jump to the next orbit.

I will start with the energy of the individuals orbits, one I will call n_f the other I will call n_i .

ΔE means the change in energy. Note that ΔE is still an energy! The change in energy from 10 J to 2 J is 8 J, which is still energy. Thus, we can use $\Delta E = h\nu$ if we need it.

$$\begin{aligned}\Delta E &= E_f - E_i \\ &= -hcR_h\left(\frac{1}{n_f^2}\right) - \left(-hcR_h\left(\frac{1}{n_i^2}\right)\right) \\ &= -hcR_h\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right) \quad (\text{we factor the } -hcR_h \text{ out})\end{aligned}$$

This is commonly written as $\Delta E = -2.179 \times 10^{-18} \text{ J} \left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right)$ since

$$-hcR_h = -2.179 \times 10^{-18} \text{ J}$$

We can also write the equation as $h\nu = -hcR_h\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right)$. This allows us to find

the frequency of the photon that must be inserted/emitted to allow the electron to jump between orbits.

The equation for hydrogen like species is not shown but you simply add Z squared.

Special note about this part.. When a question says light of a certain wavelength/frequency is emitted, ΔE is negative. This means that even though they give you a wavelength (which is a positive value), you must add the negative sign or you will get the wrong answer. Absorption has positive ΔE . The way I like to think is that emission is losing energy. Thus, you put a negative sign in front of the energy value to signify that you are subtracting that energy away. Similarly, absorption is gaining energy. Thus, you put a + sign in front to signify that it is gaining energy. (+ sign is often not used because we indicate positive values by not giving it a negative sign)

Ionization energy means removing an electron from its ground state. (Which is usually $n = 1$ but they might specify other values on the midterm) Thus, your n_i would be ground state and your n_f would be ∞ (infinity).

$$\Delta E = -2.179 \times 10^{-18} \text{ J} \left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right)$$

$$\Delta E = -2.179 \times 10^{-18} \text{ J} \left(\frac{1}{\infty^2} - \frac{1}{1}\right) \quad (1/\infty = 0, \text{ if you don't believe me, try successive}$$

large numbers. i.e. 20, 200, 2000, 200000, you will see that $1/\text{very large numbers}$ will give you values close to 0)

$$\begin{aligned}\Delta E &= -2.179 \times 10^{-18} \text{ J} (0 - 1) \\ &= 2.179 \times 10^{-18} \text{ J} \quad (\text{This value is only true for hydrogen atoms})\end{aligned}$$

This is how much energy you need to input to remove an electron from hydrogen.

That's all of 8.4.