

Explaining the Spectra of Helium and Lithium using the Rydberg formula

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The spectral lines given off by Hydrogen are well known and is simply described by the Rydberg formula. However, this only works on the hydrogen atom. If we try to describe the spectra with the Rydberg formula for helium and lithium, it fails – or does it? This a paper explores the idea that the spectra of heavier elements like helium and lithium can actually be described by just providing scaling factors to the Rydberg formula to explain the spectra given off by multi-electron atoms. This shows that the spectra is not a complex multi-body problem and that there is a very simple stair case pattern that the spectra follows.

1. Explaining the Hydrogen Spectra

When you excite hydrogen gas by passing an electric current through it, it produces a particular kind of light which is composed of only certain frequencies. See:

<http://hyperphysics.phy-astr.gsu.edu/hbase/tables/hydspec.html>

One of the successes of quantum mechanics is its ability to reproduce the formula for the hydrogen spectra. The Rydberg formula predicts the spectra of hydrogen as well as the QM Schrodinger equations can. However, no one can give an explanation for why this formula works. A description of the Rydberg formula can be found at:

https://en.wikipedia.org/wiki/Rydberg_formula

It can be expressed as: $\nu = 1/\lambda = R(1/n_1^2 - 1/n_2^2)$

In this formula λ = wavelength, ν = frequency and R is the Rydberg constant. The values n_1 and n_2 represent any 2 integer values representing the electron energy levels. Using this formula, you can precisely calculate the frequencies that are in the light produced by hydrogen.

Conventional quantum mechanics explain this as the result of an electron moving from one “electron shell” to another. The problem with this picture is how do you explain why the electrons maintain these “shells”? In atoms with dozens of electrons flying about, why don’t they collide and scatter? What could possibly hold them in their proper places? Quantum mechanics does not explain this and I think that it is impossible for the electrons to maintain any “shells”.



Do we seriously think that for large atoms that electrons can maintain these stately orbits? For an atom like hydrogen, how is it this simple atom “grows” electron shells when it only has a single electron to begin with? There is simply nothing structural outside of a hydrogen nuclei that can support up to 8 external “shells” out of the nothingness of space.

So if the electrons aren’t floating around the nucleus, then where are they normally? The radical answer is that the electrons are actually just statically bound into the nucleus. This means the electrons are just stuck to the protons just like two magnets might stick together. The electron doesn’t “orbit”, nor is it in a probabilistic cloud. It is just stuck right on the proton and is generally motionless. So a hydrogen atom is just an electron stuck on a proton. I have a whole new description of the atom which is based upon this concept called the Cubic Atomic Model. It is described by this paper:

<http://vixra.org/pdf/1303.0184v1.pdf>

The cubic atomic model does not assume that the electrons are outside of the nucleus. It assumes that they are an integral part of the nucleus. So how is the spectra of hydrogen generated if there are no electron shells to generate the spectra?

The way I derive this is by assuming that space is quantized, that is, space is made out of fixed sized grains, like the grains of sand in a beach. Let’s say we call L , the diameter of the grain of space. This restricts the movements of electrons to only move in whole integer distances $n \cdot L$ from the nucleus of the atom.

When you excite an atom, electrons are knocked loose from the nucleus and they can only travel integer distances away from the nucleus. So the main reason why electrons appear to have specific energy levels is because they can only exist at specific distances away from the nucleus. So there is no need to postulate that these electrons somehow exist in a mysterious “energy shell” floating around the nucleus with no apparent support. Instead when an atom is in the ground state, all the electrons fall straight into the nucleus and stop.

To derive the Rydberg formula using this quantized space model, we can trivially calculate the force between a hydrogen nucleus and an electron using Coulomb's law:

$F = KQ_1Q_2/R^2$ where $r = n*L$ and Q_1, Q_2 are the charges of a proton and an electron.

Since we're dealing with hydrogen with a single +1 charge and -1 charge, it simplifies to:

$$F = K/R^2$$

I claim that the energy of the spectral lines emitted is exactly proportional to the difference in electrostatic force as calculated by Coulomb's law for any 2 values of r where $R = n*L$.

So for a transition from 2 to 3, we calculate the difference. For simplification, we take $L=1$ (the diameter of an ether particle) in arbitrary units. Since we are only talking about proportionality, the constant K can be anything, so let us set it to $K=1$ so it drops out as well:

$$\text{Difference in force for a } 2 \rightarrow 3 \text{ transition} = 1/2^2 - 1/3^2$$

If we were to take a different transition, say 3 to 5, we get:

$$\text{Difference in force} = 1/3^2 - 1/5^2$$

In general, we have:

Difference in force = $1/n_1^2 - 1/n_2^2$ where n_1 is the starting level and n_2 is the finishing level. Now note that this is EXACTLY, the same format as the Rydberg formula!

Since I claim proportionality, we can add in any constant, so we can add the Rydberg constant R and equate it directly to energy.

$$\text{Energy} = R(1/n_1^2 - 1/n_2^2)$$

And since energy is inversely proportional to wavelength for light, we can write the complete Rydberg formula as:

$$1/\text{wavelength} = R(1/n_1^2 - 1/n_2^2)$$

This interpretation immediately solves some questions about the Rydberg formula. The use of $1/r^2$ terms is not just happenstance, it is a directly result of the forces calculated by Coulomb's law. It also answers how an electron should actually be perceived when it is radiating. It is simply an electron, moving fixed distances towards and away from the central proton as it being ionized. When it is not ionized, the electron falls to the lowest level basically resting on the proton. The quantization is not due to some magical something that forces the electron into fixed sized orbits, it is space itself which does the quantization. This is a concept, which I think is much easier to accept. Johann Balmer who discovered this formula could not justify how electrons maintain their stately orbits around the nucleus, but this derivation solves that problem by forcing the electron to only exist at fixed distanc-

es away from the nucleus for short periods of time due to the granular nature of space itself.

For the hydrogen atom, we can experimentally observe the electron being up to 8 steps away from the central hydrogen proton before it falls back down to the proton and releases the kinetic energy as a photon. This is like dropping a bowling ball on the ground. The higher up you lift the ball, the bigger crash it is going to make when it hits the floor. If you were limited to lifting the ball in only 1 foot increments, the crash you would get would fall into very specific energies. This is exactly what we observe in the hydrogen spectra.

2. Calculating the spectra of Helium

It should be noted that use of the QM Schrodinger equations to calculate the spectra of Helium or any other element is a complete failure. No simple formula or equations exist and the problem has to be treated as an ultra complex multi-body problem that can only be approximated on a super computer. This is because you have to consider not only the electron that is transitioning orbits, but you need to consider all the other electrons that are supposedly floating around outside the nucleus.

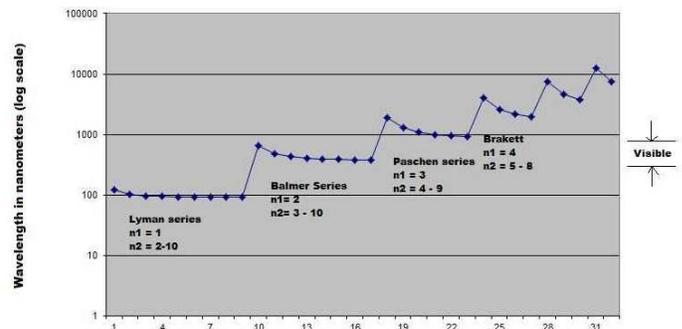
However, if you adopt the idea that the electrons which are not participating in the transition are safely locked away inside the atomic nucleus, then you can get back to the very simple picture of a single electron stepping outside of a central charge.

The spectra of Helium appears very complicated compared to hydrogen

The list of spectral data can be found at:

http://physics.nist.gov/PhysRefData/ASD/lines_form.html

We have seen that the Rydberg formula can be used to calculate the atomic spectra of Hydrogen. What we need to do is to find a "pattern" to the data. If we just group the spectral data so that we group all the lines that belong to the same n_1 together, the graph reveals a very interesting staircase pattern.



In this graph, the Lyman series shows all the transitions that start with $n_1 = 1$. The first datapoint on the left shows the transition from $n_1 = 1$ to $n_2 = 2$. It goes to $n_2 = 10$. It can be seen that this corresponds to the shortest wavelength, highest energy transitions. We then see the Balmer series starting at $n_1 = 2$. Despite the obvious pattern seen in the data, there are no references to any-

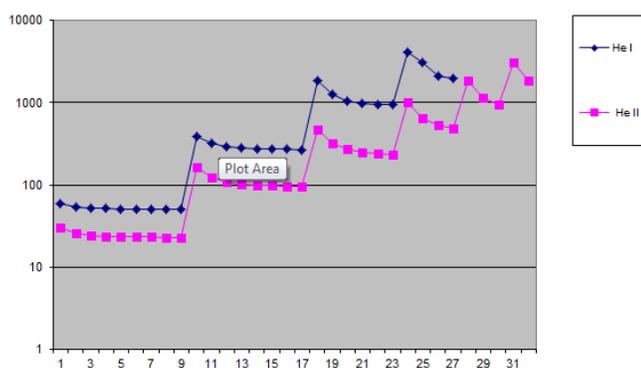
one actually making a plot like this to show this very specific pattern.

The mainstream also recognizes that the Rydberg formula can be extended to calculate the spectra of Hydrogen-like ions by adding a factor of Z^2 . These ions are any atoms which have been fully ionized and a single electron interacts with only the positively charged nucleus. For example, if you remove the 2 electrons from Helium, this ion is identified as He II. If you only remove 1 electron, this is called He I. The spectra for He II can be accurately calculated using only the extended Rydberg formula.

$$1/\lambda = R Z^2(1/n_1^2 - 1/n_2^2)$$

If you plot the calculated He II spectra wavelength against the electron level transitions (e.g. 1->2, 1->3, ... 2->3, 2->4..etc.) on a logarithmic scale, the graph appears like a staircase leading up which looks just like the hydrogen spectra. So what might start out as looking as a random mess, actually follows a very strict pattern. When the starting energy level increase by 1 (e.g 1->2 vs. 2->3), this causes a large step up in the frequency. The calculated spectra matches up well with the observed He II spectra.

Using the data from the NIST database the spectral lines for He II and He I can be graphed. Remember that the He II lines are completely predicted by the extended Rydberg formula. But we can clearly see that the He I spectra also appears as this same staircase pattern, but at a longer wavelength. This nearly identical shape suggests that the He I spectra can be calculated as simply being a scaled version of the He II spectra.



The Excel spreadsheet used to create this graph and a Microsoft word document of this article is available upon request (spectra-HE.xls).

The NIST data contains the electronic transitions and the ions that represent the observed wavelength. In most cases, this is what was used to plot the point on the graph. The Y axis shows the wavelength in nanometers. The X axis lists the 32 electron transitions in ascending order.

Since the He I spectrum appears so close to the He II spectrum, is there a simple formula that can transform one into the other? This is basically a curve fitting exercise. It does appear that a different formula can be applied at each energy level to calculate the spectral wavelengths.

For $N1 = 1$ (transitions starting at $N1 = 1$), the curve can be represented by the formula:

$$\text{Rydberg}(N1,N2)+28.14-(N2)*0.044$$

Where $\text{Rydberg}(N1,N2)$ represents the result of the extended Rydberg formula for He II and $N1$ represents the starting electron level and $N2$ represents the ending electron level. For example, the 1->2 transition calculates to 58.427 nm and matches with the observed result of 58.43 nm. This appears to be a constant minus a scaling factor based on $N2$.

For $N1 = 2$, the formula changes to:

$$\text{Rydberg}(1,N2)*13-N2*3.55$$

This appears to be a scaling factor on the Rydberg formula minus a scaling factor of $N2$, however, the shape of the curve better matches $N1=1$, so the Rydberg formula is shown with a starting level $N1=1$.

For $n = 3,4,5,6$ the formula is represented as:

$$\text{Rydberg}(N1,N2)*4$$

This is a straight scaling by 4 for all of the higher energy level shells.

The following is the calculated He I spectral data based on these formulas. Columns $N1$ & $N2$ represent the electron shell transition number. The third column represents the observed wavelength and fourth represents the calculated wavelength. The last column shows the difference between the observed and calculated values which is typically less than a percent difference. If you were to plot the calculated values over the experimental values, you could hardly tell the difference.

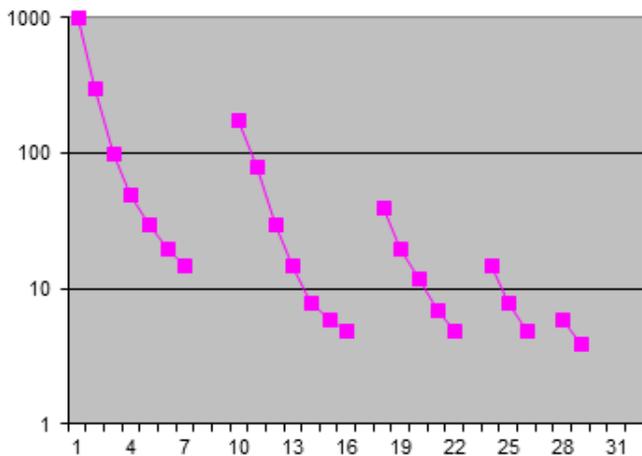
$N1$	$N2$	Observe	Calc	% Difference
1	2	58.43	58.43	0.00
1	3	53.70	53.64	0.12
1	4	52.22	52.26	0.09
1	5	51.56	51.65	0.18
1	6	51.20	51.31	0.21
1	7	50.99	51.09	0.19
1	8	50.86	50.93	0.14
1	9	50.77	50.81	0.08
1	10	50.70	50.71	0.02
2	3	388.87	387.78	0.28
2	4	318.77	322.53	1.17
2	5	294.51	301.71	2.38
2	6	282.91	290.75	2.70
2	7	276.38	283.32	2.45
2	8	272.32	277.48	1.86
2	9	269.61	272.46	1.05
2	10	267.71	267.91	0.08
3	4	1868.53	1874.61	0.32
3	5	1279.06	1281.47	0.19
3	6	1031.12	1093.52	5.71
3	7	970.26	1004.67	3.43
3	8	960.34	954.35	0.63
3	9	952.62	922.66	3.25

4	5	4048.99	4050.08	0.03
4	6	3091.69	2624.45	17.80
4	7	2113.78	2164.95	2.36
4	8	1954.31	1944.04	0.53
5	6		7455.82	
5	7		4651.26	
5	8		3738.53	
6	7		12365.19	
6	8		7498.43	

These calculations do not account for all of the 96 observed He I spectral lines. This limited calculation provides at most 32 values and for He I, there are no observed values for transitions starting from the 5 & 6 level, so this calculation accounts for 27 of 96 lines. There may certainly be other processes involved which create the other observed values, however, it is remarkable how closely the observed values can be matched with the calculated ones.

Another aspect of atomic spectra which appears to be absent in the literature is the calculation of the relative intensity for the spectral lines. If you do a similar plot of energy level transitions against the relative intensity found in the NIST data for hydrogen this produces a very regular saw tooth shape where each group for similar n1 drops down rapidly.

Hydrogen



The relative intensity can be calculated for N1=1,2,3 as:

$$(1/N1^3 * 1000) * 1 / (N2 - N1)^{(2.23 - (N1^2) * 0.13)}$$

For N1 = 4,5,6

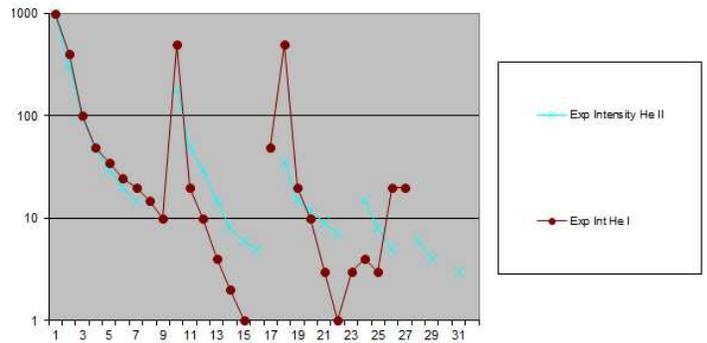
$$(1/N1^3 * 1000) * 1 / (N2 - N1)$$

The intensity appears to drop as the inverse cube. These calculations are able to reproduce the observed relative intensities as found in the NIST data.

This exact same formula also appears to apply to the relative intensity for He II. The graph for Helium shows the experimentally observed values in light blue He II. For He I, the intensity follows the formula for N1=1, it partially follows when N1=2, but

after that the relative intensity becomes chaotic and does not appear to follow any pattern.

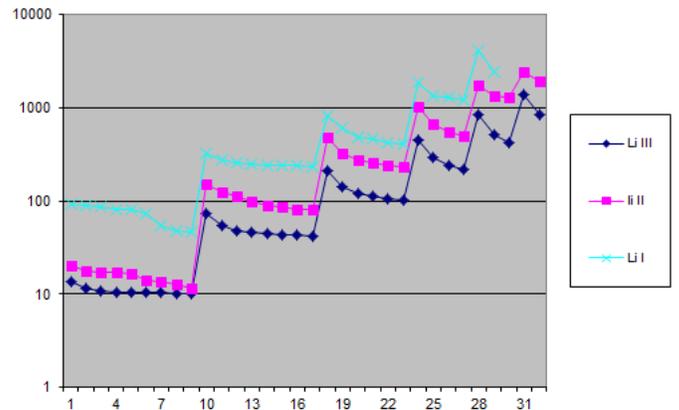
Helium



3. Calculating the spectra of Lithium

The same spectra analysis can be applied to the next element in the periodic table which is lithium. The graph of the lithium looks like this:

Lithium



This is just a simple plot of the line data found in the NIST database ordered by species and grouped by transitions. What may have appeared to be a random scatter chart of lines, now appears as a very ordered set of data.

For lithium II and I, the electron transition states stated in the NIST data do not directly provide all of the transition states required to plot the points on this graph. In this case, some of the points have been selected based upon where one would "predict" where a point would exist on the graph and using those points having the greatest relative intensity. The regularity of the pattern allows you to make these predictions. If you take a ruler and match up with any 2 of the peaks of the steps, you can predict where the next step should appear. After this peak, one would expect to find a set of $1/N2^2$ decreasing values and relative intensities. By using this methodology, the experimental data has been matched to the electron transitions as shown in pink in the graph.

The Li II values (shown as yellow in the graph) can be calculated with the following formulas:

For $N1=1$
 $Rydberg(N1,N2)+6.42$

For $N1=2$
 $Rydberg(N1,N2)*2$

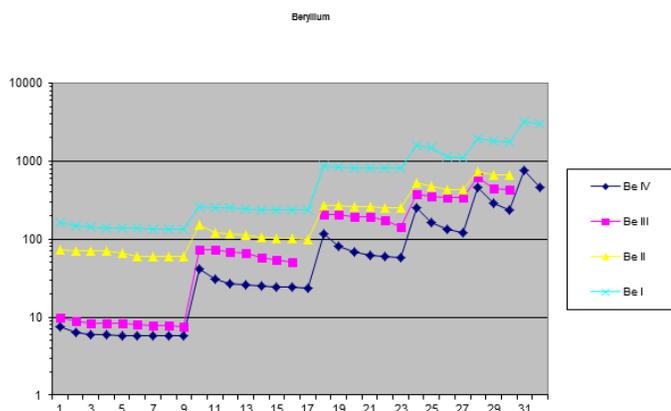
For $N1=3,4,6$
 $Rydberg(N1,N2)*2.3$

For $N1=5$
 $Rydberg(N1,N2)*2.145+(N2-5)*124$

Spectral data points can also be found for Li I which has 3 electrons bound to the atom with only 1 free electron. This also shows the familiar staircase pattern. The formulas for Li I have not been calculated, but it should be obvious that a best fit formula could be found for these data points.

The calculations match fairly closely with the observed Li data except for $N1=1$ and $N2=5-9$. Here we see an unusual dip in the wavelength compared to calculations. This dip can also be seen in the Li I data as well, so there may be something structural in Li that causes this deviation.

The analysis for the next element Beryllium becomes even more difficult as we are faced with a nearly continuous spectra. For example, to do the analysis of Be I, the only points that appeared to be known with any confidence are the very highest transitions which have very few spectral lines in the longest wavelengths. Based on the slope of those data points, the rest of the points were selected. For Be, even the ion identification needed to be occasionally ignored to find best fit data points. Due to this difficulty, the exact quality of the data point selection is questionable. However, there is enough data to speculate that the overall shape and slope of the graph is correct and the graph contains most of the brightest observed lines.



Compared with lithium, we see that the Li III and Li II are closely spaced like Be IV and Be III. The next ion Li I and Be II appear to be further spaced away. Be I appears closely spaced with Be II. This similar pattern suggests that the ions may follow a predicta-

ble energy pattern and it provides confidence that the Be graph correctly describes the energy pattern.

4. Conclusion

From the spectral data for hydrogen through beryllium, a regular pattern can be seen in the data. The spectra may initially appear to be a random collection of oddly spaced lines, but when you look at the pattern of overlapping wavelengths created by the different ions, it is easy to see how such a pattern is created.

A relatively crude curve matching formula was created to match this pattern which was based only upon the starting and ending electron level $N1$ & $N2$. It is possible that more sophisticated analysis will reveal an even simpler formula to describe the very regular staircase pattern found in the data.

Since the spectra can be described entirely as a function of $N1$ & $N2$, it would appear that the problem of calculating spectra may not be a complex multi-body problem as was previously thought. Whatever effect that the electrons have in shielding the nucleus, this effect appears to be constant and so you only need to consider the nucleus and the electron as a simple two body problem like it is in the original Rydberg formula.

However, not all of the spectra can be explained, as there are still numerous unexplained lines. However, by eliminating the points which can be explained, it may be possible to find further patterns in the unexplained lines. These other lines may require complex multi-body calculations. These calculations also do not take into account any fine differences such as the lamb shift. It only covers transitions that can be described using $N1$ and $N2$ as parameters.

Only the first four elements have been examined using this analysis. This can be extended to the other elements as well and other regularities may appear which may further enhance our understanding of the atom. The formulas derived could also have other uses such as in the generation of synthetic spectra for use in astronomy and may lead to a more accurate understanding of which lines belong to what electronic transition. These formulas may allow the prediction and detection of as of yet undiscovered spectral lines.

It is extremely surprising that the regularity of the spectral lines has not been prominently noted in the literature. The analysis done here is extremely simple and obvious. The regularity of the H and He II relative spectral intensity should be part of any standard description of the spectra for H and He as it follows a very regular pattern.

The formulas presented may have been created Ad Hoc to match the data and it is unclear why they have the form that they do. However, like Bohr and Rydberg who were also unable to explain why the spectra appear the way they do, it is important to continue to explore and find patterns within the data that can be described by simple formulas in the hopes that one will find the underlying mechanisms. This analysis deserves further research and may open new avenues in the science of the atom.