8.2 Atomic Spectra
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<u>Key equations:</u>
Balmer's Equation (I will show you how to derive this in 8.4, for now, just know this.

$\nu = 3.2881$ x 10^{15} $s^{-1}(\frac{1}{2^2} - \frac{1}{n^2})$, where n >2, (i.e. n = 3, 4, 5, 6…)

Ok, we have just learned about EMR in 8.2. Lets see how does atomic spectra connect with EMR. Once you see how everything flows together in chemistry, it is simply beautiful.

Keep in mind that light at this stage is still assumed to be a EMR wave.

First of all, lets answer the question, how is atomic spectra created? Here is the experiment that produced atomic spectra.

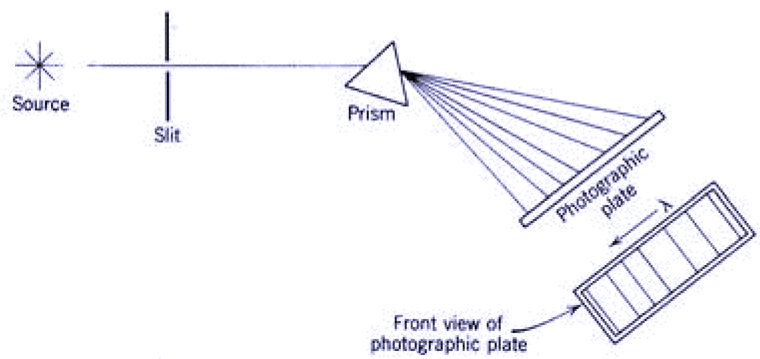
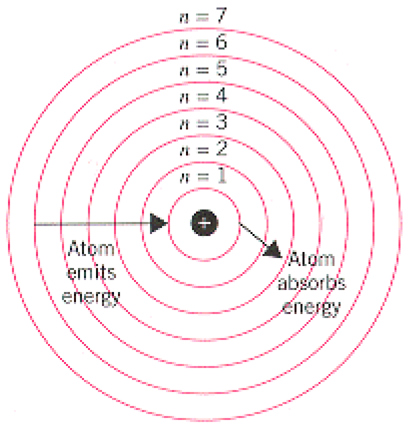


Figure 5–1. An apparatus used to measure atomic spectra.

We have the source, which is really an element that has been vaporized (converted to the gas state. The reason we vaporized elements is because that it is much easier to bump electrons within an atom if the element is in a gaseous state). This element (hydrogen gas was used in the first experiment performed in the 1800s) is stored in this glass tube. Now, I am going to pass an electric current through the gas tube. This electric current has energy,

Now, these hydrogen atoms inside the tube takes the incoming energy. Their electrons take this energy and they are going to do something with in. We have not talked about energy orbits before so lets start now. As you are probably aware, atoms have energy orbits inside them. (We will talk about orbitals later; for now, just think in terms of orbits) The idea is that each higher up orbit have a higher energy associated with them . (You probably know this from the reading of 8.4 that you have done, we will talk about this in detail in 8.4)



So, to reach a higher orbit, electrons would need an input of energy from the electric current. Electrons are going to use the energy they gained from the EMR and jump up. Now read this closely. Atoms with excited electrons are not stable. Thus, these atoms will discard their excess energy ASAP by giving off energy in the form of electromagnetic radiation.

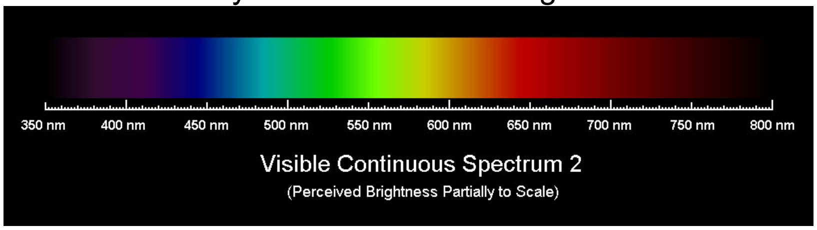
Now, you are probably confused.. Energy first came from electric current and then released as EMR??? How does this happen? To be honest, I have no idea why. I could search an answer for it but it will be a waste of time as it is not something you need to know for the midterm. Key idea is, energy comes, atom absorbs them, this excite their electrons and they jump up to higher orbits. They jump down immediately because atoms with excited electrons are not stable. When they jump down, EMR is released.

I hope you followed along. Now, we are going to place a prism to disperse the EMR into different wavelengths. The EMR released has wavelengths in the visible spectrum, the UV spectrum, and the IR spectrum. The reason for this is that a tube of say hydrogen gas has many many hydrogen atoms. They will all absorb energy but sometimes atoms bump into each other and lose a significant amount of energy. (Think of a ball, when it is moving, it has a lot of energy in the form of kinetic energy – energy possessed simply because it is moving. When the ball bumps into a wall and stops, it loses all energy. Electrons do not lose all of their energy, their energy is simply reduced) So, what happens is that electrons that possess more energy will jump off from higher orbits than those with lower energy. Thus, you get a wide range of frequencies of EMRs that's been released. Make sense so far?

We can only observe the visible spectrum with our naked eyes. Thus, the prism disperses the EMR in the visible spectrum into different wavelengths. (Think of the incoming EMR as a number of EMRs bunched together, the prism is an object that separates them) We project what we get from the prism into a black screen and viola.. we get something!

You know that lights of different wavelengths have different colors right? Good, because you see different colors on the black screen.

This is before the quantization of energy stuff, so at that time period, people assume that they would see something like this on their screen…



BUT, something else happened!!
Instead of the continuous spectrum that they are suppose to observe, they see this!

All the chemists/physicists starts to panic! What is this?!?! They can't explain it! Why are only certain wavelengths of EMR been released? Does this means that an atom only takes in certain energies? (hence only release wavelengths of the energy that they took in) This doesn't fit in with their theory! They have always assumed in their theory that energy is continuous, which means that any energy will work (hence all wavelengths of the EMR will be observed) Something is definitely wrong! We will continue with this idea in section 8.3. For now, just know that this is one of the three experiments that showed that energy is quantized (a fancy term that means that only certain energy are allowed) and the only experiment that also showed that energy is quantized within an atom.

Lets go over some details about the atomic spectra now…

You see the discontinuous spectrum with the lines right? Those lines are discrete (meaning they are separate from each other, stand alone) It is also observed that different elements have different discontinuous spectrum. This is very interesting hmm…

Then Balmer came along. Balmer, highschool teacher, didn't know much. He observed hydrogen's discontinuous spectrum for a long time. Balmer, smart guy, good with math. He observed four lines in the discontinuous spectrum. All right, he said. Lets make equation that would fit the positions of each line. So he came up with an equation. He had no idea how it works. He had his equation and that's it.

His equation is simply…

$E = -hcR_h\,((\frac{1}{2^2}-\frac{1}{n^2})$, where E is negative (it is emission)

But since $E = h\nu$, thus negative $E = -h\nu$ (we will learn it in 8.4..)

$-h\nu = -hcR_h(\frac{1}{2^2}-\frac{1}{n^2})$

Cancel h and negative sign from two sides out..

ν = $cR_h(\frac{1}{2^2}-\frac{1}{n^2})$ you now have this equation.. how is this different from the equation at the first page? Is the same!

At the very top of 1st page, you had…

$\nu = 3.2881$ x 10^{15} $s^{-1}(\frac{1}{2^2}-\frac{1}{n^2})$

If you multiply c by R_h, you get 3.2881 x 10^{15}, isn't that neat?

The value for n is any integer >2 because if you get a n that is smaller than 2, you would get a negative answer for the frequency, which is impossible.. if you have n=2, you would get 0 as the frequency.

Ah! I almost forget, I mentioned that Balmer used his equation to fit into the lines he observed. How is he suppose to do that when his equation calculates frequency and the lines are positioned by their wavelengths?

Remember the relationship between wavelength and frequency? Let me remind you..

$$\upsilon = \frac{c}{\lambda}$$

ok, so we sub this into the equation…

$$\frac{c}{\lambda} = cR_h(\frac{1}{2^2}-\frac{1}{n^2})$$

Look, the 2 c (s) cancel!

$$\frac{1}{\lambda} = R_h(\frac{1}{2^2}-\frac{1}{n^2})$$

There you have it, the wavelength equation. (well is 1/wavelength but you know how to calculate the wavelength from here right?)

So you if you sub in n = 3, 4, 5, 6, you should be able to calculate all the wavelengths of the emission line for hydrogen in the visible spectrum.

You might be wondering.. what if n > 6? Well there are only 4 lines in the hydrogen spectrum.. where does the 5th line go? A jump from n = 7 to n= 2 release more energy than a jump from n = 5 to n =2 right ? Thus, anything higher than n =6 would release EMRs that have their wavelengths in the UV spectrum. (Remember, UV have higher energy than the visible!)

Well, 8.2 doesn't have a lot! 8.3 is coming next and is huge!