5.111 Principles of Chemical Science, Fall 2008

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Transcript – Lecture 5

PROFESSOR: All right. As you're settling in, why don't you take 10 more seconds to answer the clicker question. This is the last question we'll see in class on the photoelectric effect, so hopefully we can have a very high success rate here to show we are all ready to move on with our lives here.

OK, good. So, most of you did get the answer correct. For those of you that didn't, you, of course, can ask your TA's about this in recitation, they'll always have a copy of these slides. But just to point out the confusion can be we've actually switched what the question is here. What the information we gave was the work function, which is what we've been giving before, but now we gave you the kinetic energy of the ejected electron, so you just need to rearrange your equation so now you're solving for the incoming energy, which would mean that you need to add those two energies together. So, hopefully everyone that didn't get this right, can look at it again and think about asking it's just asking the question in a different kind of a way.

All right. So, we can switch over to the class notes. So today, we're going to start talking about the hydrogen atom. Now that we have our Schrodinger equation for the hydrogen atom, we can talk about it very specifically in terms of binding energies and also in terms of orbitals. And we talked about on Wednesday the conditions that allowed us to use quantum mechanics, which then enable us to have the Schrodinger equation, which we can apply, and part of that is the wave particle duality of light and matter.

So there was a good question in Wednesday's class about the de Broglie wavelength and if it can actually go to infinity. So I just wanted to address that quickly before we move on, and actually address another thing about dealing with wavelengths of particles that sometimes comes up.

So, the question that we had in last class, was if we have actually a macroscopic particle, and the velocity let's say starts to approach zero, shouldn't we have the wavelength go to infinity, even if we have a magic board, and even if the mass is really large. So, in most cases, you would think that as the velocity gets very tiny, the mass is still going to be large enough to cancel it out and still make it such that the wavelength is going to be pretty small, right. Because if we think about the h, Planck's constant, here that's measured in 10 to the negative 34 joules per second. So, we would actually need a really, really, really tiny velocity here to actually overcome the size of the mass, if we're talking about macroscopic particles, to have a wavelength that's going to be on the order. So, let's say we're talking about the baseball, have a wavelength of the baseball that's on the order of the baseball.

So, if we kind of think about the numbers we would need, we would actually need a velocity that approached something that's about 10 to the negative 30 meters per second. So first of all, that's pretty slow here. It's going to be hard to measure anyway, and in fact, if we're talking about something going 10 to the negative 30, and we're going to observe it using our eyes, so you using visible light to observe something going this slow, we're actually not going to be able to do it because we're limited by the wavelength of light to see how precise we can measure where the actual position is. So let's say we have the wavelength of light somewhere on the order of 10 to the negative 5 meters, being the wavelength of light, we're only going to be able to measure the velocity because of the uncertainty principle to a certain degree. And it turns out it's three orders of magnitude that -- the uncertainty is three orders of magnitude bigger than the velocity that we're actually trying to observe to get to a point where we could see the wavelength, for example, for even a baseball that is moving this slow?

So that the more complete answer to the question is that no, we're never going to be able to observe that because of the uncertainty principle it's not possible to observe a velocity that's this slow for a macroscopic object. So, hopefully that kind of clears up that question. And, of course, when the velocity actually is zero, this equation that the de Broglie has put forth is valid for anything that has momentum, so if something does not have any velocity at all, it actually does not have momentum, so you can't apply that equation anyway.

And another thing that came up, and it came up in remembering as I was writing your problem set for this week, which will be posted sometime this afternoon, your problem set 2, is when we're talking about wavelengths of particles, and for specifically for electrons sometimes, you're asked to calculate what the energy is. And I just want to remind everyone, so this is a separate thought here, that we often use the energy where energy is equal to h c divided by wavelength. So if we're talking about, for example, we know the wavelength of an electron and we're trying to find the energy or vice versa, is this an equation we can use to do that? What do you think? No. Hopefully you're going to say no.

And the reason is, and this will come up on the problems and a lot of students end up using this equation, which is why I want to head it off and mention it ahead of time, we can't use an equation because this equation is very specific for light. We know it's very specific for light because in this equation is c, the speed of light. So any time you go to use this equation, if you're trying to use it for an electron, just ask yourself first, does an electron travel at the speed of light? And if your answer is no, your answer will be no, then you just know you can't use this equation here. So instead you'd have to maybe if you start with wavelength, go over there, and then figure out velocity and do something more like kinetic energy equals 1/2 n b squared to get there. So this is just a heads up for as you start your next problem set.

All right. So jumping in to having established that, yes, particles have wave-like behavior, even though no, they're not actually photons, we can't use that equation. But we can use equations that describe waves to describe matter, and that's what we're going to be doing today. We're going to be looking at the solutions to the Schrodinger equation for a hydrogen atom, and specifically we'll be looking at the binding energy of the electron to the nucleus. So we'll be looking at the solution to this part of the Schrodinger equation where we're finding e. Then we'll go on, after we've made all of our predictions for what the energy should be, we can actually confirm whether or not we're correct, and we'll do this by looking at photon emission and photon absorption for hydrogen atoms, and we'll actually do a demo with that, too, so we can confirm it ourselves, as well as matching it with the observation of others.

And if we have time, we'll move on also to talking about the other part of the solution to the Schrodinger equation, which is psi or this wave function here. And remember I said that wave function is just a representation of the particle, particularly when we're talking about electrons -- we're familiar with the term orbitals. psi is just a description of the orbital. So, we'll start with an introduction to that if we got to it at the end.

So, to remind you, when we look at the Schrodinger equation here, we have two parts to it, so when we solve the Schrodinger equation, we're either finding psi, which as I said, is a wave function or an orbital. And in addition to finding psi, we can also solve to find e or to find the energy for any given psi, and these are the binding energies of the electron to the nucleus.

And the most important thing about using the Schrodinger equation and getting out our solutions for potential orbitals and potential energies for an electron with the nucleus, is that what we find is that quantum mechanics, and quantum mechanics allowing us to get to the Schrodinger equation, allows us to correctly predict and confirm our observations for what we can actually measure are indeed the energy levels. Here we're talking about a hydrogen atom and that's what we'll focus on today. And it's incredibly precise and we're able to make the predictions and match them with experiment. Also, when we're looking at the Schrodinger equation, it allows us to explain a stable hydrogen atom, which is something that classical mechanics did not allow us to do.

So here's the solution for a hydrogen atom, where we have the e term here is equal to everything written in green. We've got a lot of constants in this solution to the hydrogen atom, and we know what most of these mean. But remember that this whole term in green here is what is going to be equal to that binding energy between the nucleus of a hydrogen atom and the electron.

So, let's go ahead and define our variables here, they should be familiar to us. We have the mass, first of all, m is equal to m e, so that's the electron mass. We also have e, which is going to be the charge on the electron. In addition to that, we have that epsilon nought value, remember that's the permittivity constant in a vacuum, and basically that is what we use as a conversion factor to get from units. We don't want namely coulombs to units, we want that will allow us to cancel out in this equation. And finally we have Planck's constant here, which we're all familiar with.

So, what actually happens when people work with the solution to the Schrodinger equation for a hydrogen atom is that they don't always want to deal with all these constants here, so we can actually group them together and use them as a single new constant, and this new constant is the Rydberg constant. And the Rydberg constant is actually equal to 2 . 1 8 times 10 to the negative 18 joules. So when we pull out all of those constants and instead use the Rydberg constant, what it allows us to do is really simplify our energy equation. So now we have that energy is equal to the negative of the Rydberg constant divided by n squared.

So, what we have left in our equation is only one part that we haven't explained yet, and that is that n value. And it turns out that when you solve the Schrodinger equation, you find that there are only certain allowed values of this integer n. And those allowed values range anywhere from n equals 1, you can have n equal that 2, 3, and it goes all the way up to infinity. But the important part is that there are only certain allowed values, so for example, you can't have 1 . 5 or 2 . 3, there are only these interger numbers. And this n here is what we call the principle quantum number. And what we find is when we apply n and plug it in to our energy equation, is that what we see is now we don't just have one distinct answer, we don't just have one possible binding energy of the electron to the nucleus. We're going to find that we actually have a whole bunch of possible, in fact, an infinite number of possible energy levels, and that's easier to see on this energy diagram here.

So, let's start with n equals 1, since that's, of course, the simplest case. So, if we have n equals 1, we can plug it into our energy equation here, and find that the binding energy, the e sub n, for n equals 1, it's just going to be equal to the negative of the Rydberg constant, so we can actually graph that on an energy diagram here, and it's going to be down low at the bottom because that's going to be, in fact, the lowest or most negative energy when n equals 1.

But we saw from our equation that there's more than just one possible value for n, so we could, for example, have n equals 2, n equals 3, all the way up to n equaling infinity. So what this tells us here is that this is not necessarily the binding energy of the electron in a hydrogen atom, it's also possible that it could, for example, have this energy, it could have this energy up here, it could have some energy way up here. So we have this infinite number of possible binding energies. But the really important point here is that they're quantized. So it's not a continuum of energy that we can have, it's only these punctuated points of energy that are possible.

So as I tried to say on the board, we can have n equals 1, but since we can't have n equals 1/2, we actually can't have a binding energy that's anywhere in between these levels that are indicated here. And that's a really important point for something that comes out of solving the Schrodinger equation is this quantization of energy levels.

And thanks to our equation simplified here, it's very easy for us to figure out what actually the allowed energy levels are. So for n equals 2, what would the binding energy be? Someone shout it out.

Yup. So, I think the compilation of the voices that I heard was negative r h over 2 squared. We can do the same thing for 3, negative r h over 3 squared is going to be our binding energy. For 4, we can go all the way up to infinity, and actually when we get to the point where it's infinity, what we find is the binding energy at that point is going to be zero. And when we get to infinity, what that means is that we now have a free electron, so now the electron has totally separated from the atom. And that makes sense because we're at the point where there's no binding energy keeping it stable.

You'll also know that all of these binding energies here are negative, so the negative sign indicates that it's low. It's a more negative energy, it's a lower energy state. So whenever we're thinking about energy states, it's always more stable to be more low in an energy well, so that's why it makes sense that it's favorable, in fact, to have an electron interacting with the nucleus that stabilizes and lowers the energy of that electron by doing so.

So, we actually term this n equals 1 state gets a special name, which we call the ground state, and it's called the ground state because it is, in fact, the lowest to the ground that we can get. It's the most negative and most stable energy level that we have. And when we think about kind of in a more out practical sense what we mean by all of these binding energies, another way that we can put it is to give it some physical significance, and the physical significance of binding energies is that they're equal to the negative of the ionization energies.

So, for example, in a hydrogen atom, if you take the binding energy, the negative of that is going to be how much energy you have to put in to ionize the hydrogen atom. So, if, for example, we were looking at a hydrogen atom in the case where we have the n equals 1 state, so the electron is in that ground state, the ionization energy, it makes sense, is going to be the difference between the ground state and the energy it takes to be a free electron.

When we graph that on our chart here, it becomes clear that yes, in fact, the ionization energy is just the negative of the binding energy, so we can just look over here and figure out what our ionization energy is. So when we're talking about the ground state of a hydrogen atom, our ionization energy is just the negative of the Rydberg constant, so that easy, it's 2 . 1 8 times 10 to the negative 18 joules.

So, that should make a lot of sense intuitively, because it makes sense that if we need to ionize an atom, we need to put energy into the atom in order to eject that electron, and that energy we need to put in better be the difference between where we are now and where we have to be to be a free electron. So in most cases when we talk about ionization energy, if we don't say anything specific to the state we're talking about, you should always assume that we are, in fact, talking about the ground state.

So, oftentimes you'll just be asked about ionization energy. If it doesn't say anything else we do mean n equals 1. But, in fact, we can also talk about the ionization energy of different states of the hydrogen atom or of any atom. So, for example, we could talk about the n equals 2 state, so that's this state here, and it's also what we could call the first excited state. So we have the ground state, and if we excite an electron into the next closest state, we're at the first excited state, or the n equals 2 state.

So, we can now calculate the ionization energy here. it's an easy calculation -- we're just taking the negative of the binding energy, again that makes sense, because it's this difference in energy here. So what we get is that the binding energy, when it's negative, the ionization energy is 5 . 4 5 times 10 to the negative 19 joules.

So we should be able to think about these binding energies and figure out the ionization energy for any state that were asked about. So if we can switch over to a clicker question here and we'll let you do that. And what we're asking you to do is now tell us what the ionization energy is of a hydrogen atom that is in its third excited state.

All right. Let's take 10 more seconds on that. OK. Interesting. Usually the majority is correct, but actually what you did was illustrate a point that I really wanted to stress and there's no better way to stress it then to get it incorrect, especially when it doesn't count, it doesn't hurt so bad. So, if you want to switch back over to the notes, we'll explain why, in fact, the correct answer is 4. So the key word here is that we asked you to identify the third excited state. So, what white is n equal to for the third excited state? 4 OK.

So that explains probably most of the confusion here and you just want to be careful when you're reading the problems that that's what you read correctly. I think everyone would now get the clicker question correct. So, the third excited state, is n equal to 4, because n equals 2 is first excited, 3 is second excited, 4 is third excited state. So now we can just take the negative of that binding energy here, and I've just rounded up here or 1 . 4 times 10 to the negative 19 joules.

So, I noticed that a few, a very, very small proportion of you, did type in selections that were negative ionization energies. And I'll just say it right now you can absolutely never have a negative ionization energy, so that's good to remember as well. And intuitively, it should make sense, right, because ionization energy is the amount of energy you need to put in to eject an electron from an atom. So you don't want to put in a negative energy, that's not going to help you out, you need to put in positive energy to get an electron out of the system. So that's why you'll find binding energies are always negative, and ionization energies are always going to be positive, or you could look at the equation and see it from there as well.

All right. So, using the equation we'd initially discussed, the negative r sub h over n squared, we could figure out all of the different ionization energies and binding energies for a hydrogen atom, and it turns out if we change the equation only slightly to add a negative z squared in there, so, negative z squared times the Rydberg constant over n squared, now let's us calculate energy levels for absolutely any atom as long as this one important stipulation, it only has 1 electron in it. So basically we're dealing with hydrogen atoms and then we're going to be dealing with ions.

So, for example, a helium plus 1 ion has 1 electron at z equal 2. A lithium 2 plus ion has 1 electron, it has z equals 3, so if we were to plug in, we would just do z squared up here, or 3 squared. Terbium 64 plus, another 1 electron atom. What is z for that? Yup. 65. So again, terbium 64 plus, not an ion we probably will run into, but if we did, we could, in fact, calculate all of the energy levels for it using this equation here.

And the difference between the equation, the reason that that z squared comes in there is because if you go back to your notes from Wednesday, and you look at the long written out form of the Schrodinger equation for a hydrogen atom, or any 1 electron atom, you see the last term there is a coulomb potential energy between the electron and the nucleus. So, of course, when we have a charge on the nucleus equal to 1, as we do in a hydrogen atom, the z is equal to 1, so it drops out there, but normally we would have to include the full charge on the nucleus, which is equal to z or the atomic number times the electron. So even if we strip an atom of all of its electrons, we still have that same amount of positive charge in the nucleus.

So, this allows us to look at a bunch of different atoms, of course, limited to the fact that it has to be a 1 electron atom.

So, now that we can calculate the binding energies, we can think about is this, in fact, what matches up with what's been observed, or, in fact, could we predict what we will observe in different kinds of situations now that we know how to use the binding energy, and hopefully we can and we will. So one thing we could do is we could look at the different wavelengths of light that are emitted by hydrogen atoms that are excited to a higher state. So what we'll do in a few minutes here is try this with hydrogen. So we'll take h 2 and we'll run -- or actually we'll have h 2 filled in an evacuated glass tube. When we increase the potential between the 2 electrodes that we have in the tube -- we actually split the h 2 into the individual hydrogen atoms, and not only do that, but also excite the atoms.

So when you just run across an atom in the street, you can assume it's going to be in its most stable ground state, that's where the electron would be, but when we add energy to the system, we can actually excite it up into all different sorts of higher states -- n equals 6, n equals 10, any of those higher states. But that only happens momentarily, because, of course, if you have an energy in a higher energy level, it's going to want to drop back down to that lower or more stable level. And when it does that it's going to give off some energy equal to the difference between those two levels. And that will be associated with a wavelength if it releases the energy in terms of a photon.

So that's what we'll look at in a few minutes. There's some important things to point out about what is happening here. Just to visualize exactly what we're saying, what we're saying is when we have an energy in a higher energy level, so let's say energy level, this initial level high up here, and it drops down to a lower final level, what we find is that the photon that is going to be emitted is going to be emitted with the exact energy, and the important term here is the exact. That is the difference between these two energy states.

That makes sense because we're losing energy, we're going to a level lower level, so we can give off that extra in the form of light. And we can actually write the equation for what we would expect the energy for the light to be. So this delta energy here is very simply the energy of the initial state minus the energy of the final state.

This is a little bit generic, we're not actually specifying the states here, but we could, we know we can calculate the energy from any of the states. So, for example, let's say we excited the hydrogen atom such that the electron was starting in the n equals 6 state, so that's our n initial. And we drop down to the n equals 2 state, or the first excited state. Then we would be able to change our equation to make it a little bit more specific and say that delta energy here is equal to energy of n equals 6, minus the energy of the n equals 2 state.

When we talk about that frequency of light that's going to be emitted, it's not too commonly that we'll actually talk about it in terms of energy. A lot of times we talk about light in terms of its frequency or it's wavelength, but that's OK, because we know how to convert from energy to frequency, so we can do that here as well, where our frequency is just our energy divided by Planck's constant, and since here we're talking about a delta energy, we're going to talk about the frequency as being equal to delta energy over h. Or we can say the frequency is going to be equal to the energy initial minus the energy final all over Planck's constant.

So this means that we can go directly from the energy between two levels to the frequency of the photon that's emitted when you go between those levels. What we can also do is say something about the wavelength as well because we know the relationship between energy and frequency and wavelength. So, in the first case here, let's say we go from a high level to a low level, let's say we go from five to one. If we have a large energy difference here, are we going to have a high or low frequency? Good. A high frequency. If we have a high frequency, what about the wavelength, long or short? All right. Good. So we should always be able to keep these relationships in mind. So, similarly in a case where instead we have a small energy difference, we're going to have a low frequency, which means that we're going to have a long wavelength here.

So now we can go ahead and try observing some of this ourselves. So what we're actually going to do is this experiment that I explained here where we're going to excite hydrogen atoms such that they're electrons go into these higher energy levels, and then we're going to see if we can actually see individual wavelengths that come out of that that correspond with the energy difference.

So, our detection devices are a little bit limited here today, we're actually only going to be using our eyes, so that means that we need to stick with the visible range of the electromagnetic spectrum. Actually that simplifies things, because that really cuts down on the number of wavelengths that we're going to be trying to observe here. So it turns out that in the visible range, when you figure out the differences between energy levels, in hydrogen atoms, there's only 4 wavelengths that fall in the visible range of the spectrum. So hopefully, when we turn out the lights, we're going to turn on this lamp here, which has hydrogen in it and we're going to excite that hydrogen. You'll see light coming out, but it's, of course, going to be bulk light -- you're not going to be able to tell the individual wavelengths.

So what our TAs, actually if they can come down now, are going to pass out to you is these either defraction goggles, or just a little plate, and you're going to be splitting that light into its individual wavelengths. And the glasses, there aren't enough to go around for all of you, so that's why there's plates. And though glasses do look way cooler, the plates work a little better, so either you or your neighbor should try to have one of the plates in case one of you can't see all the lines. So our TAs will pass these around for us.

And I also want to point out, it's guaranteed pretty much you'll be able to see these three here in the visible range -- you may or may not be able to see, sometimes it's hard to see that one that's getting near the UV end of our visible spectrum. So we won't worry if we can't see that.

I'll take one actually. Can you raise your hand if you if you still need one? All right, so TA's walk carefully now, I'm going to shut the lights down here. All right, so we do still have some little bits of ambient light, so you might see a slight amount of the full continuum. But if you look through your plate, and actually especially if you kind of look off to the side, hopefully you'll be able to see the individual lines of the spectrum. Is everyone seeing that? Yeah, pretty much. OK. Can anyone not see it? Does anyone need -- actually I can't even tell if you raise your hand. So ask your neighbor if you can't see it and get one of the plates if you're having trouble seeing with the glasses. So this should match up with the spectrum that we saw. And actually keep those glasses with you. We'll turn the light back on for a second here. And hydrogen atom is what we're learning about, so that's the most relevant here. But just to show you that each atom does have its own set of spectral lines, just for fun we'll look at neon also so you can have a comparison point.

All right. So, this is probably a familiar color having seen many neon lights around everywhere. So you see with the neon is there's just a lot more lines in that orange part of the spectrum then compared to the hydrogen, and that's really what gives you that neon color in the neon signs. That's the true color of a neon being excited, sometimes neon signs are painted with other compounds. All right, does everyone have their fill of seeing the neon lines?

STUDENT: No.

PROFESSOR: All right, let's take two more seconds to look at neon then.

All right. So our special effects portion of the class is over. And what you see when we see it with our eye, which is all the wavelengths, of course, mixed together, is whichever those wavelengths is most intense. So, when we looked at the individual neon lines, it was the orange colors that was most intense, which is why we were seeing kind of a general orange glow with the neon, which is different from what we see with the hydrogen.

All right, so we can, in fact, observe individual lines. There's nothing more exciting to see with your glasses on, while you look nice. You can take those off if you wish to, or you can try to just be splitting the light in the room until the TAs grab your glasses, either is fine.

It turns out that we are far from the first people, although it felt exciting, we did not discover this for the first time here today. In fact, J.J. Balmer, who was a school teacher in the 1800s, was the first to describe these lines that could be seen from hydrogen. And he saw the same lines that we saw here today, and although he could not explain, even start to explain why you saw only these specific lines and not a whole continuum of the light. Right, we already have an idea because we just talked about energy levels, we know there's only certain allowed energy levels, but at the time there's no reason J.J. Balmer should have known this, and in fact he didn't, but he still came up with a quantitative way to describe what was going on. He came up with this equation here where what he found was that he could explain the wavelengths of these different lines by multiplying 1 over 4 minus 1 over some integer n, and multiplying it by this number, 3 . 29 times 10 to the 15 Hertz, and he found that this was true where n was some integer value -- 3, 4, 5, or 6.

So he could explain it quantitatively in terms of putting an equation with it, but he couldn't explain what was actually going on. But we, having learned about energy levels, having had the Schrodinger equation solved for us to understand what's going on, can, in fact, explain what happened when we saw these different colors.

So, what we know is happening is that were having transitions from some excited states to a more relaxed lower, more stable state in the hydrogen atom. And it turns out what we can detect visibly with our eyes is in the visible range, and that means that the final state is n equals 2. Because you see how n equals 1 is so much further away, and actually that's not to scale, it's actually much, much further down in the energy well, such that the energy of the light is so great that it's going to be in ultraviolet very high energy, high frequency range. So we can't actually see any of that, it's too high energy for us to see. So everything we see is going to be where we have the final energy state being n equals 2.

So if we think about, for example, this red line here, which energy state or which principle quantum number do you think that our electron started in?

STUDENT: Three.

PROFESSOR: Good. So, it's going to be in 3, because that's the shortest energy difference we can have, and the red is the longest wavelength we can see -- those 2 are inversely related, so it must be n equals 3. What about the kind of cyan-ish, blue-green one? Yup. so n equals 4 for that one. Similarly we can go on, match up the others. So n equals 5 for the bluish-purple and the violet is n equals 6. And again, that matches up, because the violet, or getting really close to the UV range here has the longest energy, so the highest frequency, and that's going to be the shortest wavelength and we can see here it is, in fact, the shortest wavelength that we can actually see.

So, we can see if we can come up with the same equation that J.J. Balmer came up with by actually starting with what we know and working our way that way instead of coming up from the other direction, which he did, which was just to explain what he saw. So, if we start instead with talking about the energy levels, we can relate these to frequency, because we already said that frequency is related to, or it's equal to the initial energy level here minus the final energy level there over Planck's constant to get us to frequency. And we also have the equation that comes out of Schrodinger equation that tell us exactly what that binding energy is, that binding energy is just equal to the negative of the Rydberg constant over n squared.

So that means that our frequency is going to equal, if we plug in e n into the initial and final energy here, 1 over Planck's constant times negative r h over n initial squared, minus negative r h over n final squared. So we have an equation that should relate how we can actually calculate the frequency to what J.J. Balmer observed. We can simplify this equation by pulling up the r h and getting rid of some of these negatives here by saying the frequency is going to be equal to the Rydberg constant divided by Planck's constant all times 1 over n final squared -- this is just to switch the signs around and get rid of some negatives -- minus 1 over n initial squared. We can plug this in further when we're talking about the visible part of the light spectrum, because we know that for n final equals 2, then that would mean we plug in 2 squared here, so what we get is 1 over 4. So this is our final equation, and this is actually called the Balmer series, which was named after Balmer, and this tells us the frequency of any of the lights where we start with an electron in some higher energy level and we drop down to an n final that's equal to 2. So it's a more specific version of the equation where we have the n final equal to 2. And it turns out that actually we find that this matches up perfectly with Balmer's equation, because the value of r h, the Rydberg constant divided by the Planck's constant is actually -- it's also another constant, so we can write it as this kind of strange looking cursive r here. Unfortunately, this is also call the Rydberg constant, so it's a little bit confusing. But really it means the Rydberg constant divided by h, and that's equal to 3 . 29 times 10 to the 15 per second. So if you remember what the equation Balmer found was this number multiplied by this here. So, we found the exact same equation, but just now starting from understanding the difference between energy levels.

So, sometimes you'll find the Rydberg constant in different forms, but just make sure you pay attention to units because then you won't mess them up, because this is in inverse seconds here, the other Rydberg constant is in joules, so you'll be able to use what you need depending on how you're using that constant.

So in talking about the hydrogen atom, they actually have different names for different series, which means in terms of different n values that we end in. So we talked about what we could see with visible light, we said that's actually the Balmer series. So anything that goes from a higher energy level to 2 is going to be falling within the Balmer series, which is in the visible range of the spectrum. We can think about the Lyman series, which is where n equals 1. We know that that's going to be a higher energy difference, so that means that we're going to be in the UV range. We can also go in the opposite direction. So, for example, when n equals 3, that's called the Paschen series, and these are named after basically the people that first discovered these different lines and characterized them, and this is in the near IR range. And again, the n equals 4 is the Brackett series, and that's an IR range. I think there's names for even a few more levels up. You don't need to know those, but just because it's a special case with the hydrogen atom, they do tend to be named -- the most important, of course, tends to be the Balmer series because that's what we can actually see being emitted from the hydrogen atom.

So, now we should be able to relate what we know about different frequencies and different wavelengths, so Darcy can you switch us over to a clicker question here? And we can also talk about the difference between what's happening when we have emission, and we're going to switch over to absorption. So, we just talked about emission, so before we head into absorption, if you can answer this clicker question in terms of what do you think absorption means having just discussed hydrogen emission here? So we have four choices in terms of initial and final energy levels, and also what it means in terms of the electron -- whether it's gaining energy or whether it's going to be emitting energy? So, why do you take 10 seconds on that, we'll make it a quick one.

All right, great. So already just from knowing the emission part, we can figure out what absorption probably means. Absorption is just the opposite of emission, so instead of starting at a high energy level and dropping down, when we absorb light we start low and we absorb energy to bring ourselves up to an n final that's higher. And instead of having the electron giving off energy as a photon, instead now the electron is going to take in energy from light and move up to that higher level.

So now we're going to be talking briefly about photon absorption here. So again, this is just stating the same thing, and it could take in a long wavelength light, which would give it just a little bit of energy, maybe just enough to head up one energy level or two, or we could take in a high energy photon, and that means that the electron is going to get to move up to an even higher energy level.

And again, we can talk about the same relationship here, so it's a very similar equation to the Rydberg equation that we saw earlier, except now what you see is the n initial and the n final are swapped places. So instead now we have r h over Planck's constant times 1 over n initial squared minus n final squared. And what you want to keep in mind is that whatever you're dealing with, whether it's absorption or emission, the frequency of the light is always going to be a positive number, so you always want to make sure what's inside these brackets here turns out to be positive. So that's just a little bit of a check for yourself, and it should make sense because what you're doing is you're calculating the difference between energy levels, so you just need to flip around which you put first to end up with a positive number here, and that's a little bit of a check that you can do what yourself.

So let's do a sample calculation now using this Rydberg formula, and we'll switch back to emission, and the reason that we'll do that is because it would be nice to actually approve what we just saw here and calculate the frequency of one of our lines in the wavelength of one of the lines we saw.

So what we'll do is this problem here, which is let's calculate out what the wavelength of radiation would be emitted from a hydrogen atom if we start at the n equals 3 level and we go down to the n equals 2 level. So what we need to do here is use the Rydberg formula, and actually you'll be given the Rydberg formulas in both forms, both or absorption and emission on the exams. if you don't want to use that, you can also derive it as we did every time, it should intuitively make sense how we got there. But the exams are pretty short, so we don't want you doing that every time, so we'll save the 2 minutes and give you the equations directly, but it's still important to know how to use them.

So, we can get from these energy differences to frequency by frequency is equal to r sub h over Planck's constant times 1 over n final squared minus 1 over n initial squared. So let's actually just simplify this to the other version of the Rydberg constant, since we can use that here. So kind of that strange cursive r, and our n final is 2, so 1 over 2 squared minus n initial, so 1 over 3 squared. So, our frequency of light is going to be equal to r times 5 over 36. But when we were actually looking at our different wavelengths, what we associate mostly with color is the wave length of the light and not the frequency of the light, so let's look at wavelength instead. We know that wavelength is equal to c over nu. We can plug in what we have for nu, so we have 36 c divided by 5, and that cursivey Rydberg constant, and that gives us 36 times the speed of light, 2 . 998 times 10 the 8 meters per second, all over 5 times 3 . 29 times 10 to the 15 per second.

So, what we end up getting when we do this calculation is the wavelength of light, which is equal to 6 . 57 times 10 to the negative 7 meters, or if we convert that to nanometers, we have 657 nanometers. So does anyone remember what range of light 657 nanometers falls in? What color?

STUDENT: Red.

PROFESSOR: Yeah, it's in the red range. So that's promising. We did, in fact, see red in our spectrum, and it turns out that that's exactly the wavelength that we see is that we're at 657 nanometers.

So it turns out that we can, in fact, use the energy levels to predict, and we could if we wanted to do them for all of the different wavelengths of light that we observed, and also all the different wavelengths of light that can be detected, even if we can't observe them.

All right, so that's what we're going to cover in terms of the energy portion of the Schrodinger equation. I mentioned that we can also solve for psi here, which is the wave function, and we're running a little short on time, so we'll start on Monday with solving for the wave function.