DJORGOVSKI: All right. So where do these spectroscopic lines come from? Atomic physics tells us where they come from. I take it that you probably learned, at least in chemistry, a little bit about energy levels of atoms and molecules. But let's go through this in a little more detail.

So electrons would change between different energy states inside atoms or molecules. Or they may be in free plasma just scattering off of other electrons or ions. If they're bound to an atom or a molecule, they can just jump from one level to another. Then those are bound-bound transitions.

If they absorb a photon of such an energy that it actually knocks them out of their atom or molecule, then that's called ionizations, bound-free. And the other way around-- if an electron is captured by positive ion, then it's recombination.

Electrons can just zip around and scatter off each other or positive ions. Or they can also be interacting directly with photons. That's called Compton scattering.

Now, what happens depends on physical conditions-- the temperature, the composition, the pressure of the gas, and so on, how has it been ionized. The simplest system is the hydrogen atom. This was how Bohr figured it out. And that's how quantum mechanics began.

And schematically, electron inside a hydrogen atom can't be any simpler. It is one proton. It is one electron. And so then they're different levels. Its principal quantum number is the number of level. The innermost is 1 and so on.

And the energy of any given level is a constant called the Rydberg constant, which is 13.6 electron volts for a simple hydrogen, divided by the square of the principle quantum number. And the lowest is called the ground state. Then there are fine and hyperfine transitions, but we need not go into this.

So notice their energy is inversely proportional to the square of the quantum number. Therefore, the difference between the two that corresponds to the energy

of the photon will be proportional to the difference between 1 over n1 squared and 1 over n2 squared. And this is what Bohr figured out.

So how do electrons go between different states? They can gain or lose energy in one of two ways. Either they can get photo excited-- that the absorbed photon can climb to higher energy level, or de-excited, they go down-- or collisionally excited or de-excited. Mechanical collisions of atoms can actually transfer kinetic energy and bump electrons from one level to another.

Both of those happen in nature. And usually, we see in astrophysical sources that there is photo ionization. There's some source of ultraviolet light or something like that that pumps up the higher levels. There is sometimes collisional de-excitation, depending on the pressure of the gas.

In hydrogen itself-- and that is the simplest case-- even back in the 19th century, people noticed that different emission absorption ions of hydrogen show regularities that were written as Rydberg's formula. And transitions to or from a given level correspond to one of those series of lines.

So the one that goes to or from the first level is called Lyman series. And it's important for cosmological studies, because objects far away so redshifted we can see the ultraviolet light. The one that's most frequently seen is Balmer series, which is to and from the second level of hydrogen atom. And then there is Paschen and Brackett and so on.

Now, think about the energy level. For hydrogen, which is just the simplest nucleus-- one proton-- they're of the order of a few electron volts. Now, what do you think would happen for heavier nucleus, say iron, with 56 protons it? Would its energy levels be higher or lower in the absolute sense?

Any guesses to that?

STUDENT: Higher.

Yes. They would be higher, of course, because there is a stronger electrostatic

force between nucleus and electron. And so the binding energy of the electron would be much higher. And it turns out that for heavier elements like that, the innermost shells correspond to energies of kilo electron volts, and not just electron volts, whereas the outermost shells of something like iron would be still in the range of electron volts, just like visible light.

And so this is what spectrum of starred hydrogen-- most of them do-- looks like. You can see this nice progression of lines that follows Balmer's formula-- that difference of the 1 over square of quantum number of minus 1 over 2 squared. And they start always with alpha. That's the longest wavelength. That's between whatever that level is-- 2, in this case, and 3. And then beta would be 4 to 2 and gamma would be 5 to 2 and so on.

And then it goes all the way to a limit, which is what happens when electron goes from infinity to that level. So that's the maximum energy jumped that it can have to get to that level or from that level.

Hydrogen is very abundant in the universe, and that's why this is actually a good example of how to study. So this is what it looks like. These are the kinds of pretty pictures that you see, but now you can understand what they mean.

You see there's a whole bunch of bright blue stars? These are regions of star formation. These stars have temperatures of 10s of thousands of degrees Kelvin and law of ultraviolet tradition. And the gas that surrounds them absorbs that light, gets excited, and then gets de-excited as time goes on. The red light is from Balmer alpha line-- 6,563 Angstroms. And the blue one tends to be from Balmer beta line at 4,800 gamma, 4,300 and on. So the colors that you see in photo ionized gas reflect wavelengths of these principal recombination lines.

Another set of astrophysical objects where you can see photo ionization in action are planetary nebulae. Planetary nebulae are results of a relatively old star shedding its envelope. What's left is very hot incandescent core, which, because it's hot, shines mostly at ultraviolet wavelengths. The gas that used to be stellar envelopes, now expanding in a vacuum of space, is being eliminated by this

radiation and excited to whatever levels, and then shins in these different spectroscopic lines.

These are enhanced colors, of course, but you can see that there are different colors of different spacings. And that would correspond to different densities of ionization field of photons. It would be bluest closest to the source and then reddest further out. There's actually an interesting little story about this, which is story of nebulium, the nonexistent chemical element.

So when first spectra of these nebulae were obtained-- sometime in the early 20th century-- people noticed that there are lines there that just did not correspond to anything in the lab, which was unusual. Everything else they can sort out.

And so the hypothesis was that this was some previously unobserved chemical element. The periodic table wasn't quite full back then. It was called nebulium. It turned out there wasn't. It was just oxygen.

But these were the transitions from ionized oxygen atoms that had lost two electrons that are very hard to sustain in the lab, because that energy state gets very easily collisionally de-excited. But in space, it's almost a perfect vacuum, so you see there's a nebula. That's just because there's so many ions there in a large volume. The spacing is so large that collisional de-excitation doesn't really play a large role. So you get to see the lines that are very difficult to reproduce in the lab. So in that sense, astrophysics provides you ways of doing physics that you can't do easily in the lab on planet Earth.

You'll see the notation. Just out of curiosity, if you're wondering where it's come from, usually, there is element-- in this case, oxygen. Then there is the Roman numeral. And that's ionization state. I is the neutral. Atoms did not lose any electrons. II, they lost one electron. III, they lost two electrons, and so on.

So neutral hydrogen is sometimes called HI. And ionized hydrogen, which means a whole bunch of loose protons and electrons, is sometimes call HII. And those are the shiny, nice nebulae that you see regions of star formation. Now, if it's one of

those so-called forbidden transitions-- they're not forbidden, they're just hard to do in a lab-- then it gets these square brackets around.

OK. What about molecules? Well, in addition to the principle quantum levels inside the atoms, molecules, since they have spatial extent, can do two other things. They can vibrate and they can rotate. Both of those would have energy levels associated with them.

And those tend to be much lower energy than those from principal quantum transitions. And therefore, they correspond to spectroscopic lines in infrared or even radio. This has been useful to study cold gas and gas where there is so much absorption that it absorbs ultraviolet and visible light, but far infrared or radio would go right through. And this is how we study regions of star formation-- the clouds that are obscuring the cores of young stars.

And finally, there is an important spectroscopic transition, which is neither from principal quantum numbers changed, nor from orientation rotation. It's from spin flip. It's a hyperfine transition. In hydrogen, electron metaphorically orbiting the proton, and their spins can either be parallel or anti parallel. They correspond to two different energy levels, which are different by just a tiny amount of energy that corresponds to photon of a wave length of 21 centimeters.

And for a given atom of hydrogen, this happens, on average, once every 11 million years. But there are so many hydrogen atoms out there that this turns out to be the dominant radio signal from interstellar medium. And we use it to study structure of Milky Way and other galaxies, measure the rotation, and so on. Because these are radio waves, they go through dust clouds that obscure light. And so when this was figured out by [INAUDIBLE] and Schlovsky and others in the 1950s, this was the first time that we can actually see through the disk of the Milky Way. Otherwise, visible light just gets obscured by dust clouds.