***We Forge Into Quantum Theory:*** Now that we’ve got into the wave nature of particles like electrons, we can explore the atom itself and try to understand how it works.

At the turn of the century, physicists were totally puzzled by the behavior of atoms and electrons. They did not seem to obey Newton’s law. Here’s what seemed to be happening.

When an electron was accelerated, it would give off light and slow down. Why?

Light from the sun and from stars was a continuous spectrum – the light ran all the way from red light of long wavelength to violet light with a short wavelength. Actually this was one of Newton’s discoveries, but he used a simple prism. Later sophisticated spectroscopes were developed which allowed a really good examination of the spectrum – they spread the colors out more. With this better view of the spectrum, it was discovered that there would be places where a certain color was missing. These appeared to be dark lines in the spectrum. These missing colors were called "Fraunhofer lines" (named after the dude what discovered them). The same thing was found when a continuous spectrum of light was sent through a sample of a gas. The light that emerged from the gas would also have some of these dark lines. These came to be called absorption spectrums. Each element had its very own emission spectrum. These absorption spectrums are really useful. We can identify elements in stars and elements in unknown substances. This is done on that CSI show all the time.





Another interesting thing was discovered. When a gas was heated to incandescence (i.e., so it glowed), the gas would give off light, but only certain wavelengths. You would see various bright lines of color – and nothing else, no continuous spectrum of light like from the sun. These came to be called emission spectrums.

For a given element, the emission lines and the absorption lines occur at the same frequency.

This is where quantum mechanics comes in. Here’s the basic idea (which was the product of Niels Bohr, Erwin Schroedinger, and Verner Heisenberg).

The atom has a minimum energy state which is called its ***ground state***. Energy levels above that are called ***excited states***. In the ground state, the electrons (and therefore the atoms) have the least amount of energy they can possibly have. Above the ground state are a series of discrete allowed energy levels for the electrons. When an atom is excited, the electrons absorb the amount of energy (and only that amount) which is equivalent to one of the allowed energy levels. The electrons instantaneously jump to one of those higher energy levels. This is a now-you-see-it-now-you-don't kind of deal. They have one energy, then ……..ZAP! …..they have a new energy level - nothing in between. When an electron makes one of these jumps, it has made a ***quantum leap***. (Have you ever watched the program of the same name?) The electron doesn’t stay in the new energy level for long, it is unstable in this higher energy level, and loses the gained energy. As it does this it falls back down to a lower allowed energy level. The energy it lost is released as a photon of electromagnetic radiation. The energy amount that is absorbed or emitted is called a ***quanta***.

The quantum theory does a good job of explaining why objects can give off light when they are heated to a high temperature. Heat a steel nail in a lab burner flame. The nail absorbs thermal energy, this causes the atoms in the metal to become excited. As they gain energy, the electrons will make quantum leaps. They will jump to higher energy levels. Once there, as we said before, the electrons aren't stable and will fall back down to a lower energy state. When they fall, they release the energy they absorbed in the form of photons of electromagnetic energy - light (although the emitted electromagnetic energy is not limited to visible light, other frequencies can also be emitted, we just don't see those). Anyway, the nail glows - it gives off light. Initially, it doesn't glow very brightly - sort of a dull red glow. But as more energy is available (the nail gets hotter) the nail glows brighter and brighter. Red light is the lowest frequency of visible light, and it appears first. (Actually, infrared light shows up first, but we can't see it.) As the metal becomes hotter and hotter, the color that is seen changes. A nail will melt at a temperature that finds it glowing yellow. The filament in a light bulb is made of tungsten, which has a very high melting point, so the filament can achieve a very high temperature. It is excited by the flow of current through it. It glows red, then orange and then yellow. It finally appears to be a yellowish white. This happens because tungsten has a lot of allowed energy states and when it is really excited you end up with a continuous spectrum of light. At lower temperatures, you only get the lower energy excited states, so the thing glows in a sort of reddish orange color.



Fluorescent bulbs work on a different but similar principle (does that make sense?). A tube is filled with a gas at low pressure (the tube is almost a vacuum with just a trace of gas, say neon). A high voltage is applied across the tube (the high voltage is developed by a step-up transformer). The gas in the tube is ionized by the electric field, which allows electrons to pass through the tube. This energy excites the gas atoms, which emit the energy as light, but mainly as ultra violet light. The emitted light photons collide with a fluorescent powder that is coated on the inside of the tube. The various atoms in the powder are excited, the electrons jump up to higher energy levels, fall back down, and emit visible light photons, but they do so at a great many wavelengths, very nearly a complete spectrum. This is the light that you see.

Remember the absorption spectrum with the Fraunhofer lines? It turned out that the wavelengths absorbed by the atoms represent the photons that have the allowed energy levels for that element. The wavelengths that are found in an emission spectra also represent the allowed energy levels. See how neat all this is? If an element has a certain set of energy levels that are available to the electrons, photons that have that energy will be absorbed. This will cause the electrons to make quantum leaps to higher energy levels. Photons that have different energy amounts will not be absorbed. When the electrons drop to lower energy states, the photons they emit are limited to only those wavelengths that correspond with the allowed energy states.

The idea that only certain energy states are allowed and others are prohibited is not intuitive. But that is what happens. We can visualize it by imagining that we have a quantity of pennies. The weight of the pennies must be a multiple of the weight of a single penny, so their weight is also quantified. A pile of pennies cannot have a weight that is not a multiple of the weight of a penny.



***Electron Position:*** From chemistry, you will have learned about the various orbitals that electrons fill. Remember , the ***s***, ***p***, ***d***, and ***f*** orbitals and the suborbitals? These were found by measuring the location of electrons over time. The electrons were found to be located in certain spaces about the nucleus 90 % of the time. The volume that the electrons stayed in for the 90 % became the shape of the suborbital. You couldn't really predict where the blame electrons were going to be - they don't follow orbits like planets do. But you could be sure that 90 % of the time, the electrons would be somewhere in this funny little cloud shape.

***Why Energy Levels:***

Niels Bohr came up with idea of the allowed energy levels, but, quite frankly, they didn’t really make much sense. Why were electrons restricted in the amount of energy they could have?

De Broglies’ wave model for electrons is what came to the rescue. Schroednger thought, “What if the electrons are moving around the nucleus as a standing wave?” This would explain why only certain energy levels were allowed.

Remember standing waves? Schrodinger developed an equation that described the behavior of electrons in all the elements. This is the famous Schrodinger wave equation. Sadly, because of the mathematical difficulty level of the thing, we won’t get to play with it.

Let’s look at an energy level diagram for an atom.

Okay, right now it doesn’t look so good. Just a bunch of lines. The bottom line represents the ground state of the atom. This is its lowest energy level, above that are the allowed energy levels. In this example there are three possible energy levels. The top line is zero energy and represent s the point where the electron has left the atom (causing the atom to be ionized – “Ladies and Gentlemen, Elvis has left the building.”)



There are several possible energy transitions. Here are a few of them:

Here we have an electron at the ground state. It can leap to any of the three excited states available to it. These possible leaps are represented by the arrows. All that is needed for the electron to make a quantum leap is the proper amount of energy. If the electron does not get exactly the right amount of energy to take it to one of the excited states, it will not do so.

Let’s imagine that the electron jumps up to the third excited state. What can it then do? Well, it can make any of the following transitions:

Here’s an example of what could happen when an electron leaps to the third excited state:



The electron jumps up to the third excited state. This represents 4.5 eV of energy (5 eV-0.5 eV). It could absorb a photon that has that amount of energy in order to make this jump. Once there, it drops down to the second excited state. This represents 0.5 eV. It would release a photon that has this amount of energy when it does this. It then drops down to the ground state. This represents 4.0 eV and another photon would be released with that amount of energy. Do you see the other energy amounts that are represented?

Photons are always released when the electron falls to the lower energy state. This is where the energy goes.

See how it works?

Of course, we don’t identify photons by the energy of the thing in eV. Usually we represent them by their wavelength.

So what would be the wavelength of a photon that caused the 4.0 eV quantum leap that started things?





This photon would not be visible light – the wavelength is too short. We’re looking at UV here.

The photons emitted? Well, one represents a quantum leap of 0.5 eV. Let’s find its wavelength.



This would definitely not be visible light – wavelength is way too long.

In a typical kind of problem, you might be told that a gas absorbs a certain wavelength and gives that one off as well as another. How would that work?