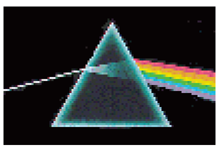
Spectral Lines

At the end of 19th century, physicists knew there were electrons inside atoms, and that the wiggling of these electrons gave off light and other electromagnetic radiation. But there was still a curious mystery to solve. Physicists would heat up different elements until they glowed, and then direct the light through a prism...



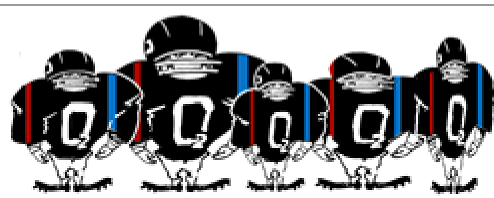
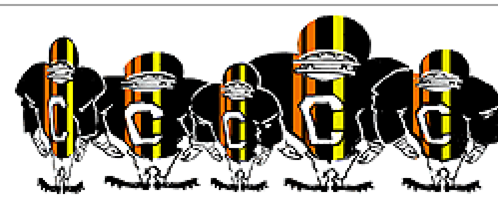
I've done that with sunlight. You see the whole rainbow because the prism breaks the light into all of its separate colors.

That's what you get with light from the sun. But when scientists looked at the light coming off of just one element, hydrogen for instance, they didn't see the whole rainbow. Instead they just got bright lines of certain colors. (Actually, "color" isn't the right term, because only some of the lines were visible, but for now we'll just talk about visible light.)

That would mean that the atoms were only emitting waves of certain frequencies. Do all atoms create the same colors?

No. Each type of atom gives off a unique set of colors. The colored lines (or **Spectral Lines**) are a kind of "signature" for the atoms.

Kind of like wearing your team colors.

 **Team Oxygen**	 **Team Carbon**

Exactly. If you put light from a common streetlamp through a prism, or look at the light through a diffraction grating, you will see distinct lines. Two common kinds of street lights use sodium vapor and mercury vapor bulbs. Each of these lights has a different spectral "signature", and you can tell what kind of lamp it is by its spectral lines.

Pick an element from the menu to see its spectral signature.

Is that why different street lights seem to be different colors?

You got it. This technique is so reliable that scientists can tell what elements they are looking at just by reading the lines. Spectroscopy (this page is currently under construction) is the science of using **spectral lines** to figure out what something is made of. That's how we know the composition of distant stars, for instance.

Wait a second. We learned earlier that radiation is caused by wiggling charges, and the rate of the wiggling determines the wavelength. If only some wavelengths are coming out of the atom, that would mean that the electrons are wiggling at only *some* frequencies. How does that happen?

That was the big puzzle. Fortunately, a Danish physicist named Niels Bohr came up with an answer...

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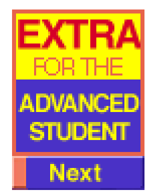
Bohr's Atom

To explain the spectral line puzzle, Bohr came up with a radical model of the atom which had electrons orbiting around a nucleus.

That doesn't sound so radical. We've already seen how electrons can orbit around a positively charged nucleus.

Yes, but in order to explain the "signature colors," Bohr came up with an extraordinary rule the electrons had to follow: **Electrons can only be in "special" orbits.** All other orbits just were not possible. They could "jump" between these special orbits, however, and when they jumped they would wiggle a little bit...



And that would cause radiation!

To see this happening, try clicking on different orbits in the model of an atom below.

Hey, when I click on a smaller orbit, a little colored squiggle goes shooting out, but when I click on a bigger orbit, a squiggle comes in and kind of "bumps" the electron up.

Those squiggles are little bursts of light (electromagnetic energy). We call them **photons**.

But when we played with the orbits earlier, we saw that just about *any* orbit and any speed is possible. It doesn't make sense that only some orbits would be "allowed."

Now you can see why the Bohr model was considered so radical! It said that energy could only change in little jumps. These are called **quanta** and that's why this kind of physics is called **Quantum Mechanics**.

Is that where the term "quantum leap" comes from?


Yup. Ironically, everyday use of the term has come to mean a *big* jump, but physicists use it to mean a jump between allowable orbits, which is usually very, *very* tiny.

The important part is that these jumps cannot be broken down into smaller steps. For an electron on the move it's all or nothing.

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Energy Levels


So that's it? Atoms really look like little solar systems with electrons making quantum jumps between special orbits?

Well, not quite. The idea of an electron actually flying around in little circles turned out to have lots of problems, and physicists were eventually forced to discard that model.


But we just finished talking about how well that worked! Why do we have to throw the whole thing away?

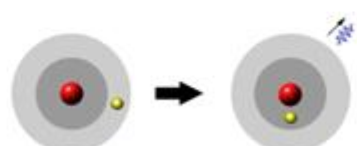
We're not going to start from scratch. The concept of "special orbits" was extremely useful, it's just the orbits themselves that we're not going to use anymore. Instead, we're going to think about electrons being in special **energy levels**. We just use this rule:

Bigger Orbit = Higher Energy

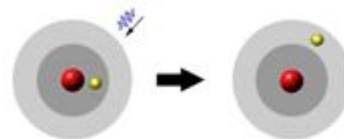
Oh, that's easy enough. But why bother? Why not just call them orbits?

Well, first of all, some orbits have the same energy as other orbits, so sometimes changing orbits wouldn't emit radiation. Also, it turns out that electrons don't *really* move in little circular orbits. We can take a little detour to see how the Schrödinger Atom more accurately depicts what is happening inside atoms.

Actually, thinking about energy levels makes more sense, anyway, because if the energy goes *down* the extra energy has to go somewhere, so it comes out as electromagnetic radiation.



Yeah, and in order for the energy to go *up* it has to come from somewhere, so it takes some incoming radiation!



This next applet shows the Bohr model along with a diagram showing the energy level. This "energy level" picture of an atom is so useful that most physicists prefer it over the orbital picture.

Hold on. Earlier we were saying that when an electron changes its speed or direction, it gives off electromagnetic radiation. Now we're saying that when an electron changes its orbit (or "energy level") it gives off electromagnetic radiation. Which is it?

You're changing your story on us! Are you making this up as you go along or what?

Change in velocity was a classical idea, but the quantum physicists realized the important part is that the *energy* of the electron changes, and electromagnetic radiation makes up the difference. If the energy goes down, the extra energy appears as a photon. And for the electron to get *more* energy, it needs to absorb a photon. Now let's look at how this theory neatly explains spectral lines...

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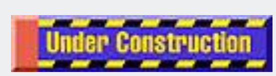
Atomic Spectra

We've seen how photons are created or destroyed when interacting with atoms, but an important thing to realize is that transitions between the same energy levels always produce the same color photon. (Actually, photons don't have colors but often that is a convenient way of thinking about their wavelength or frequency.)

I can show you how to compute the energy (and thus the frequency and wavelength) of these photons...



In the following experiment, there is a device below the Bohr model that works like a prism or a diffraction grating. It shows the atomic spectrum for a hydrogen atom. Whenever a photon is emitted, it shows up on the spectrum according to its wavelength.

Click on an orbit to make the electron jump energy levels.

Huh. Each time the electron jumps *down* a level it produces a photon, and the same jumps produce the same colors.

When you have a whole lot of atoms, I'll bet you get all these different lines appearing at once.

Exactly, and that's what scientists mean by the **atomic spectrum**. By the way, the converse is true, too. Those same color photons are the only ones that will bump the electron up to higher levels. Photons of other frequency will pass right through the atom.

That would mean atoms are kind of "transparent" to all light except their own "team colors."

We keep talking about the "color" of these photons. Does that mean that atoms only interact with visible light? What about other kinds of electromagnetic radiation?

We've been talking about visible light because it's the easiest to experiment with. But you're right, we should talk about the "frequency" or "wavelength" of the photons, not their color. In fact, we're now going to talk about how heavier atoms, which have *lots* of electrons, tend to interact with higher energy waves, like x-rays. We can go ahead and talk about these heavier atoms, or look at some specific examples, such as how hospital x-ray machines create the x-rays, or how they absorb them to make images.