

The Production of Light

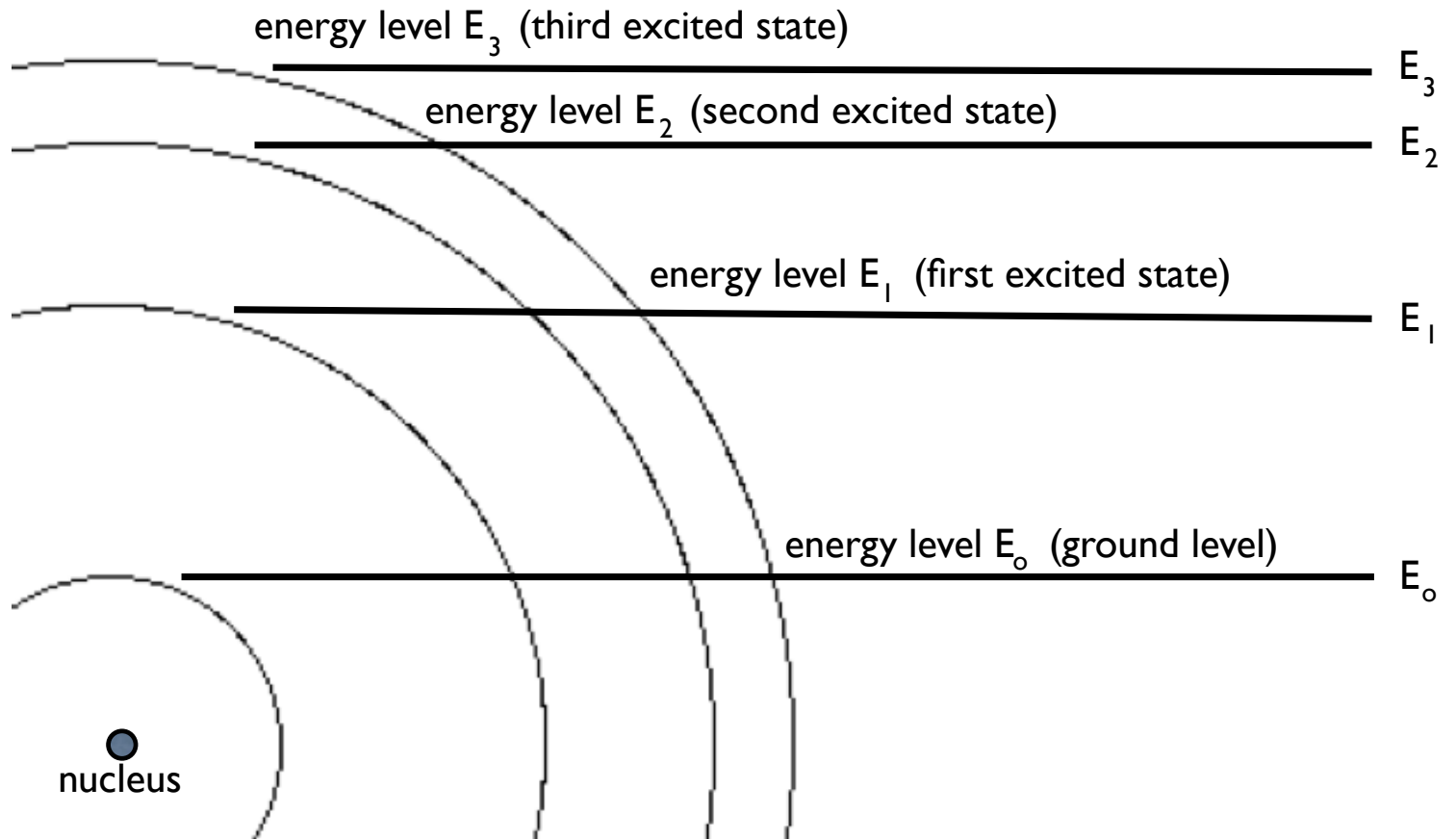
The production of light is intimately related to the way the atom is constructed.

Electrons are found in very specific areas on the periphery of the atomic structure. What determines these areas is energy—electrons can exist only in areas associated with very specific energies.

The Bohr model is a very simple construct designed to allow an observer to visualize these *orbital energy levels*.

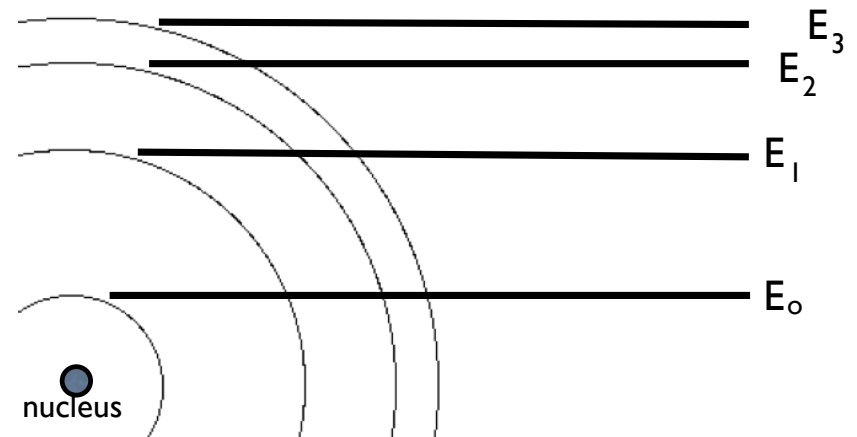
In the Bohr model, energy states are depicted by circular orbitals. The lowest possible state for an outer-shell electron (a *valence* electron) to occupy is called *the ground level*. The first state out from the ground state is called *the first excited state*, the second out *the second excited state*, etc.

The sketch to the right depicts these orbital and energy levels.



So let's say a photon of light impinges upon this atom. We know the relationship between the amount of energy E involved in that light (if viewed as a bundle of energy--a photon) and the light's frequency ν_1 (if viewed as a wave) is:

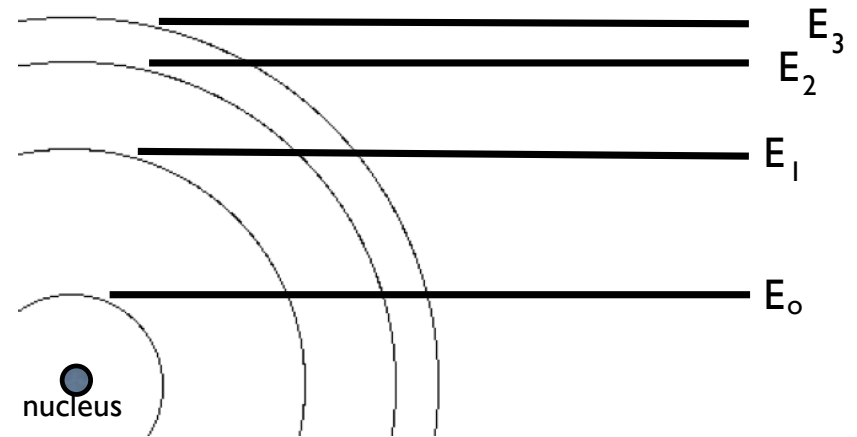
$$E = h\nu_1.$$



So what would happen if there is an electron in the ground state and that photon-energy quantity just happened to equal $E_2 - E_0$, where those E values are defined by the Bohr model (see sketch)?

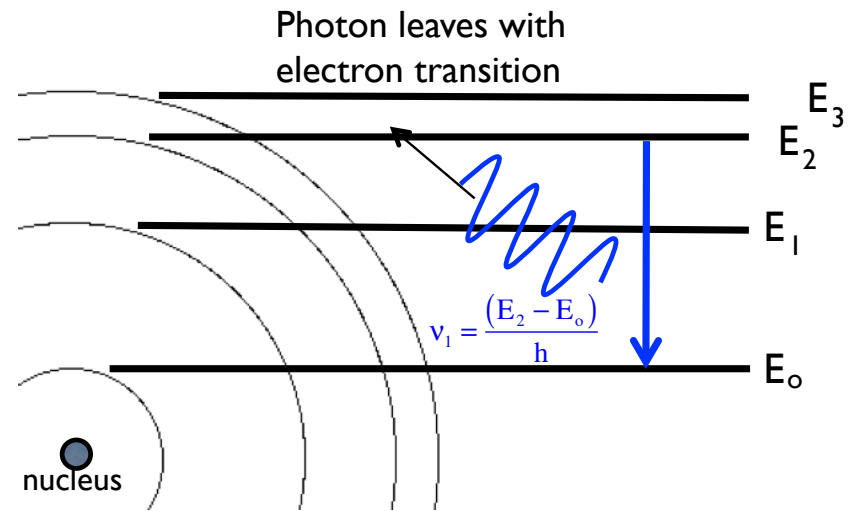
What would happen is that the atom would absorb the photon, boosting one of its electrons from the E_0 state to the E_2 state. (This would also happen if an electron just happened to be found in, say, the E_1 state and the energy content of the photons was such that it equaled $E_2 - E_1$).

If, on the other hand, the photon's energy did not match any of the energy differences between the orbit energies of the atom, then that atom would not absorb that photon (at least not in this way, though it could under the right circumstances absorb the energy by making the overall molecular structure vibrate more).

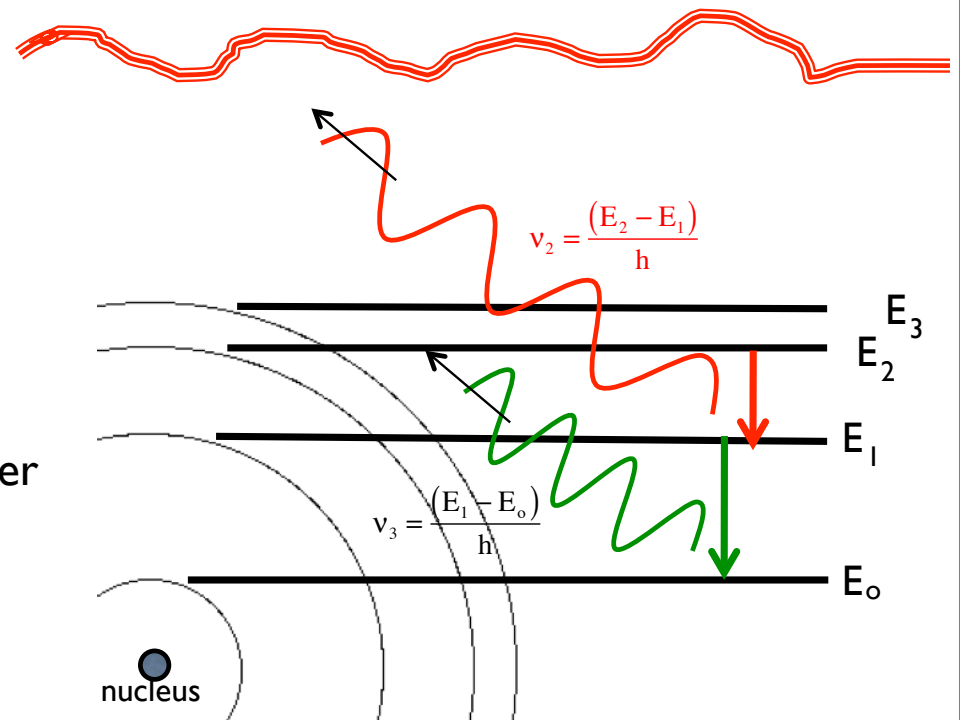


So let's assume our original scenario happens and we end up with an electron in the E_2 state. Atoms don't like being in "higher" energy levels, so in fairly short order the electron will cascade back down to the ground level. And how will that happen?

One possibility would be for the electron to dump $E_2 - E_0$ worth of energy in the form of a photon (with frequency ν_1). That would correspond to the situation sketched to right.



Another possibility is for the electron to drop from E_2 to E_1 (producing a photon of frequency ν_2), then dropping from E_1 to E_0 (producing a photon of frequency of ν_3).



In short, when an electron is excited into an upper-level energy orbital, it will sooner or later decay back down toward the ground state giving off photons whose energies correspond to the energy *differences* between the orbital levels.

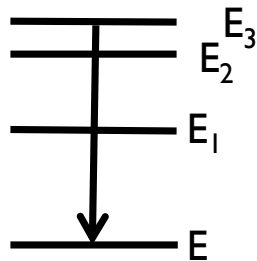
In other words, with each of these transitions, out comes electromagnetic radiation (possibly even in the optical range, or what we call “light”) of a very specific frequency.

So now we are ready for two additional points of order:

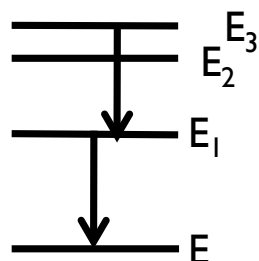
I.) **Emission spectra** and determine energy levels: If we have several billion of these atoms in, for instance, a gas, and if we excite an electron from each of those atom’s into the *third excited state* using radiation of frequency

$$\nu_6 = \frac{(E_3 - E_0)}{h},$$

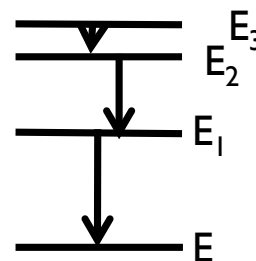
out of that gas will come a number of distinct frequencies of light, all having energies equal to one of the energy-level differences available in the atom. For this situation, they are:



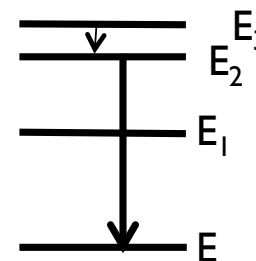
This produces one high energy, high frequency photon



This produces two more, distinct frequencies not already seen



This produces three more, distinct frequencies not already seen

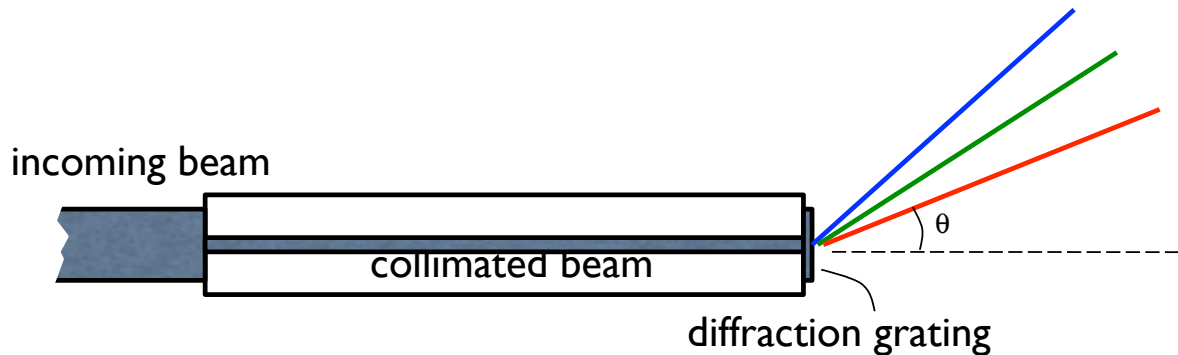


This produces one more, distinct frequencies not already seen

Enter the spectroscope.

A spectroscope is an optics device designed to analyze light from, say, a heated, glowing source. It does so by splitting the light coming from the source into the light's characteristic frequencies using a prism or diffraction grating.

A schematic of a spectroscope is shown below. The light to be analyzed travels down a collimating tube (this insure that only light rays that are parallel to one another pass through the device), striking a diffraction grating at its end (diffraction gratings are a lot like prisms in that when light passes through one, each frequency gets re-directed splitting the light into its component parts). The angle of diffraction is related mathematically to the wavelength of the light, which in turn is related to the light's frequency.



Knowing the number of *lines per centimeter* on the diffraction grating and knowing the deviation angle, one can determine the wavelength (hence frequency) of the bit of light being observed.

With a spectroscope, therefore, we can take light from any source and display the frequencies that make up that light. When we do this, what we end up with is called an *emission spectrum*.

The emission spectra for both hydrogen and helium gas is shown below.

Hydrogen



Helium



As each lit part of an emission spectra is associated with a particular wavelength, computed by that mathematical relationship alluded to earlier, each color has a calculable wavelength and, hence, frequency associated with it.

So how does that help us with energy levels? Consider the steps:

1.) Light is sent through a spectroscope with each frequency exiting at a different angle. Each angle is associated with a different, calculable wavelength and, hence, frequency. How so?

a.) The math is used on the diffraction angles to determine the actual wavelengths.

b.) The relationship $c = \lambda\nu$ is used to determine the actual frequencies.

2.) EACH OF THESE FREQUENCIES is related to an energy jump inside the atom by

$$E(= \Delta E_{\text{between orbitals}}) = h\nu$$

3.) It is possible to feed photons of different frequencies into an atom until you hit the frequency that has just the right amount of energy to liberate a ground state electron. That amount of energy, by definition, is called the *binding energy* of the atom (or, more specifically, the *ground state binding energy*). **These energy quantities are always negative as they denote how energy-deficient an electron in that state is from complete liberation.**

6.) (For **hydrogen**, for instance, that **ground state binding energy is -13.6 eV.**)

7.) The point is, if you do this you end up with ONE of the energy levels in the atom, the ground state energy.

8.) Logically, at least some of the frequencies given off by the atom will correspond to energy jumps down to the ground state, and because you know how much energy the ground state is associated with, you can deduce what the other excited states are.

Example: Let's assume we have determined that the ground-state binding energy of an atom is -13.6 eV (OK, it's a hydrogen atom—work with me). We look at all the frequencies of light that come from that gas when excited, and we decide that two of the energies are generated by electrons dropping down into the ground level. The energy changes associated with those jumps are -10.2 eV and -12.1 eV (as measured by looking at the frequency of the light being given off and using $\Delta E = h\nu$). Knowing the size of the energy jumps and knowing one energy (the ground state energy), we can determine the second energy level. That is:

$$\Delta E = E_{\text{upper},1} - E_{\text{ground}}$$

$$(-10.2 \text{ eV}) = E_{\text{upper},1} - (-13.6 \text{ eV})$$

$$\Rightarrow E_{\text{upper},1} = -3.4 \text{ eV} \quad (\text{this happens to be the first excited state of hydrogen})$$

and

$$\Delta E = E_{\text{upper},2} - E_{\text{ground}}$$

$$(-12.1 \text{ eV}) = E_{\text{upper},1} - (-13.6 \text{ eV})$$

$$\Rightarrow E_{\text{upper},1} = -1.5 \text{ eV} \quad (\text{this happens to be the second excited state of hydrogen})$$

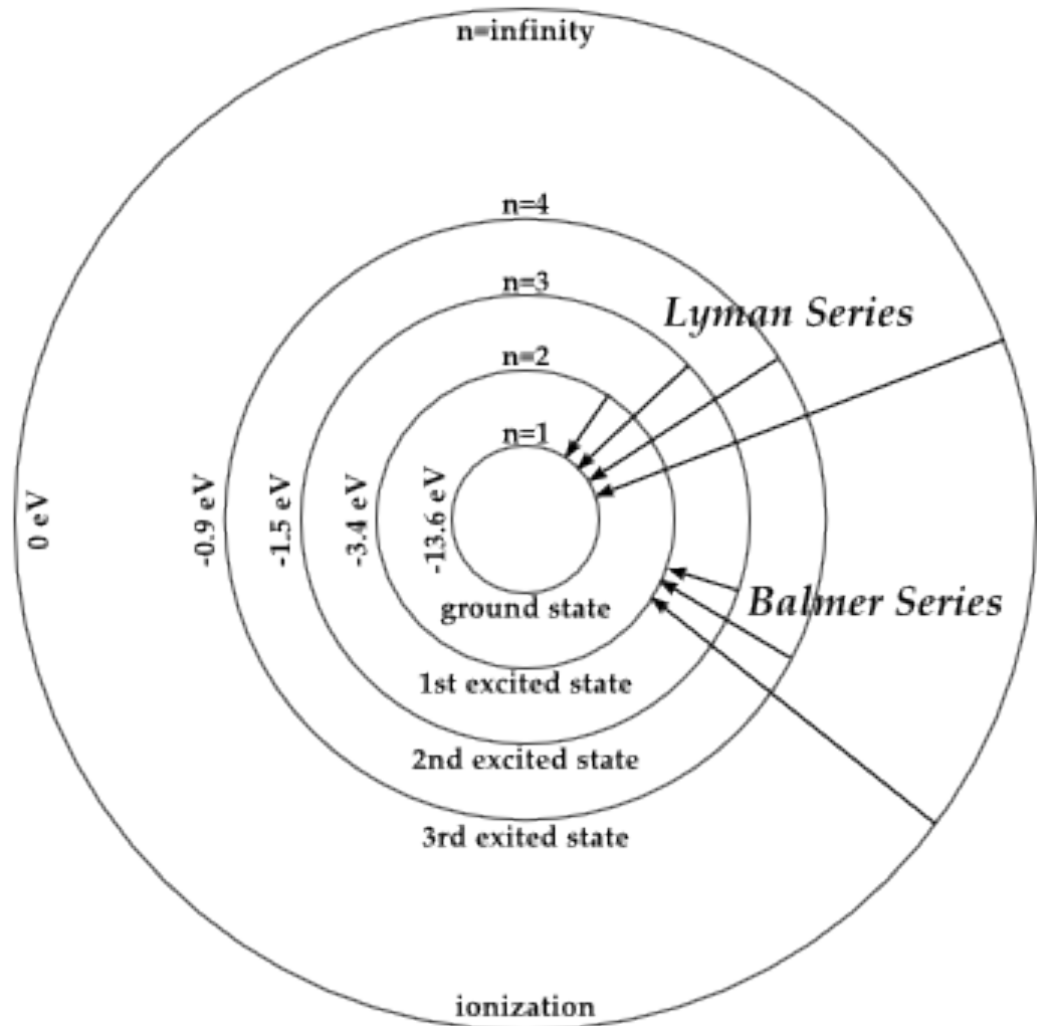
The point is that by using a spectroscope to determine the various wavelengths, hence frequencies, hence energy jumps that are generated by an element, we can determine the energies associated with each of the orbitals of that element.

NOTE:

If we had been looking at the **HYDROGEN** atom, the transitions would be as shown to the right. In that, transitions that end at the **ground state** are part of what is called the **Lyman Series**. The photons that are given off with these transitions are **not in the optical range** (in fact, they are in u.v.)

Transitions that end at the **first excited state** in hydrogen are part of what is called the **Balmer Series**. The photons that are given off with these transitions **are in the optical range**.

Transitions that end at the **second excited state** in hydrogen are part of what is called the **Paschen Series**. The photons that are given off with these transitions are generally **not in the optical range**.



(from Wikipedia)

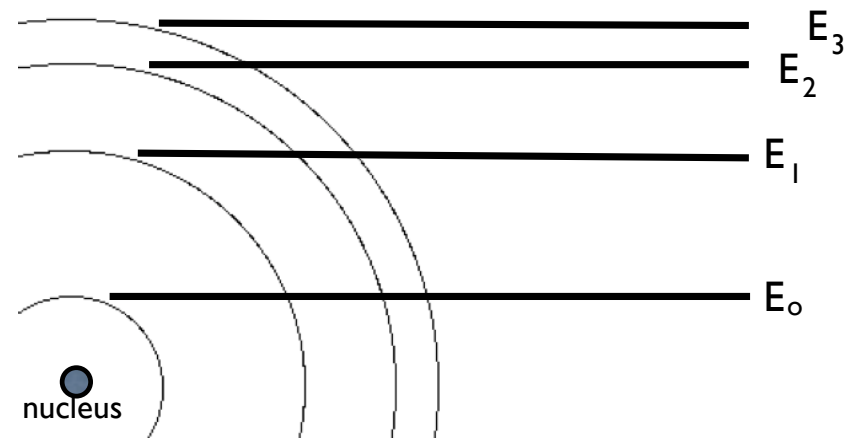
2.) **Absorption spectra:** A slightly different but related phenomenon occurs when white light (i.e., light in which all the frequencies of the optical range are present) passes through a gas.

In that case, almost all of the frequencies that pass through the gas will be unaffected and will pass straight through.

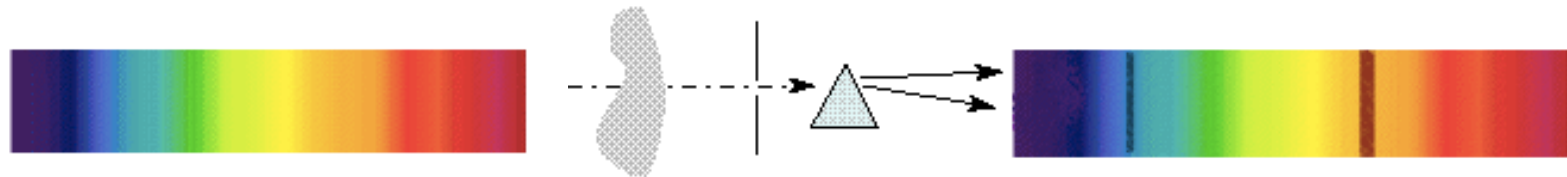
Some won't make it, though. Which ones? The ones whose energy content exactly matches the orbital energy jumps in the gas. Those frequencies will motivate electrons to make those jumps as the energy is absorbed by the atoms. And when those excited electrons decay back down to the ground level, emitting photons as they go, those photons will not necessarily be going in the direction of the original light.

In short, when the original white light, minus the absorbed frequencies, passes through a spectroscope for analysis, all the frequencies of white light will be there EXCEPT the frequencies that were absorbed out.

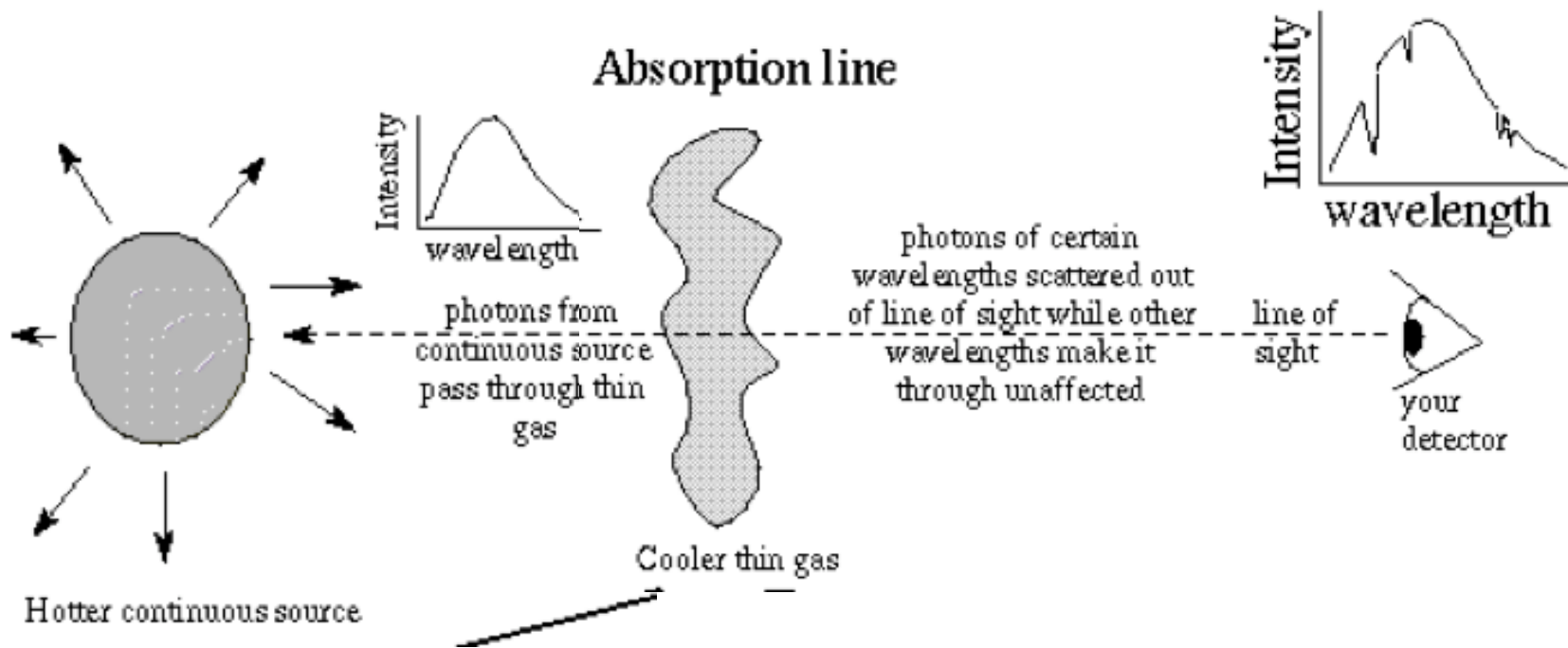
A spectrum that acts like this are called **absorption spectra**.



A pictorial representation of such a situation is shown in the schematic below.



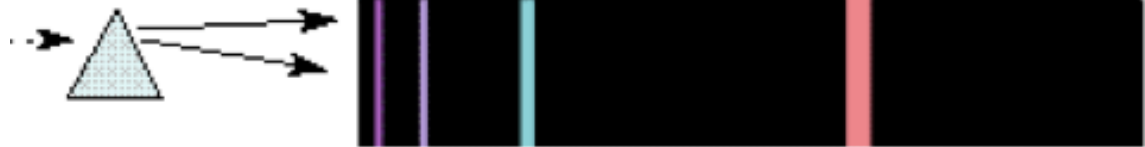
And an additional, second representation is included below:



SUMMARY

For emission spectra: Start with a specific, excited gas and send its light through a prism. On the other side of the prism, you will find very specific set of spectral lines. The frequency of each lines will be associated with the energy jump involved in the transition that produced it.

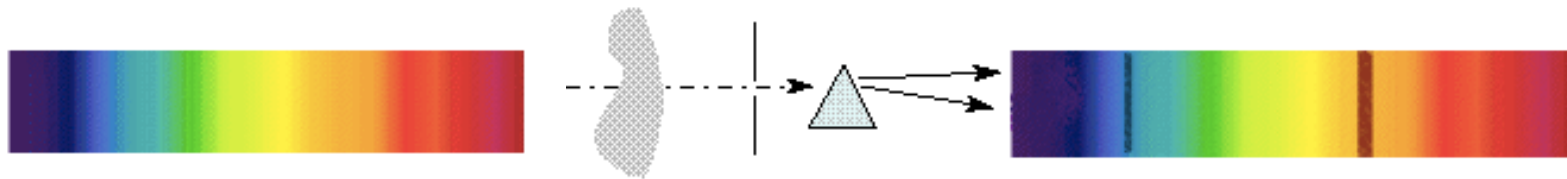
Light from an excited gas source



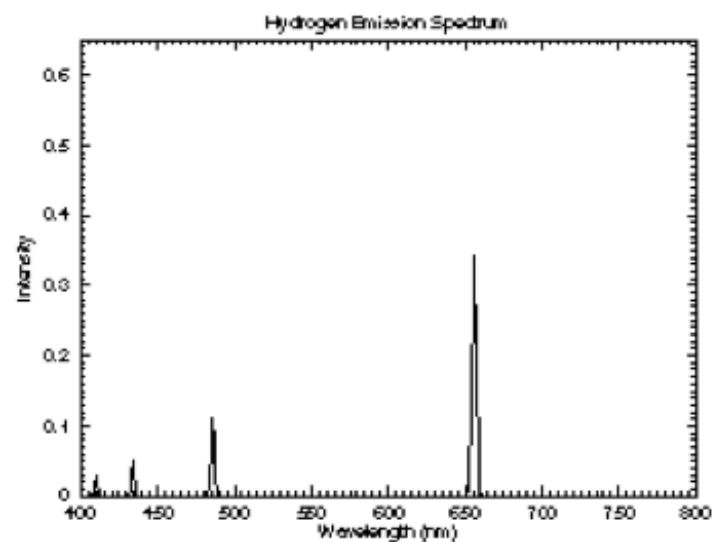
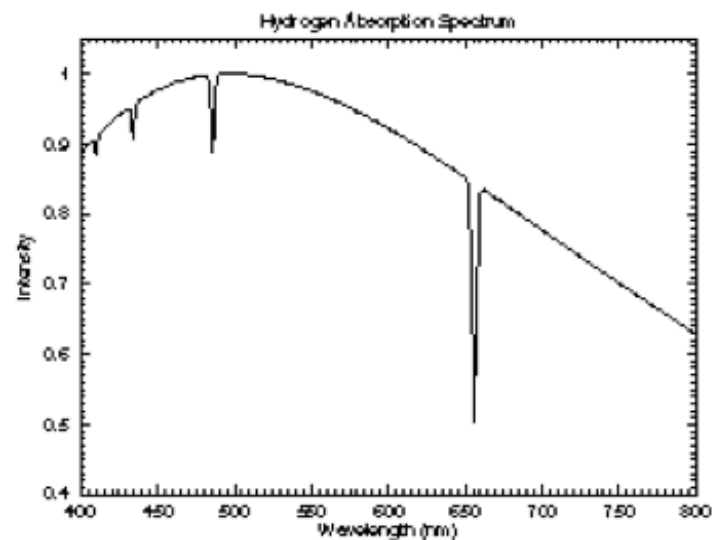
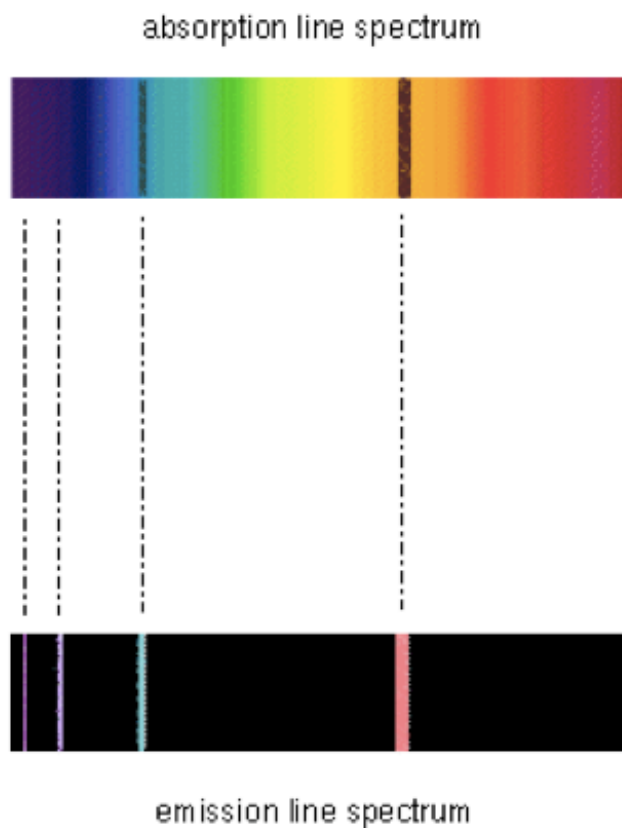
$$\text{frequency} = (E_2 - E_1)/h$$

An arrow points from the equation to the red line in the emission spectrum diagram above.

For absorption spectra: Start with a full spectrum light source. Pass the light through a specific gas. On the other side of the gas, you will find an otherwise full spectrum with specific dark lines in it. The dark lines will be due to the absorption of the frequencies of light the energies of which could excite the gas' atoms.



What is useful about all of this? Because the emission spectra for a gas and the absorption spectra for the same gas are both produced by the same energy-level differences inside the atoms of the element, the absorption lines and the emission lines should be mirror images of one another (see the example to the right).



Absorption spectra exist in any light gathered from a star (even our own star, the sun). Understanding how this spectra is produced is important . . . and this is where things get exotic.

To understand the origin of absorption spectra from a star, you need to understand something about **black body radiation**. To do that, you need to look at the PowerPoint titled, “Spectral Characteristics and Blackbody Radiation.” Go there now, then come back to this PowerPoint

2.) With the information you accumulated from the “Spectral Characteristics and Blackbody Radiation” PowerPoint, you should now understand where a star’s absorption spectrum comes from. With that, we are now ready for the *second point of order*.

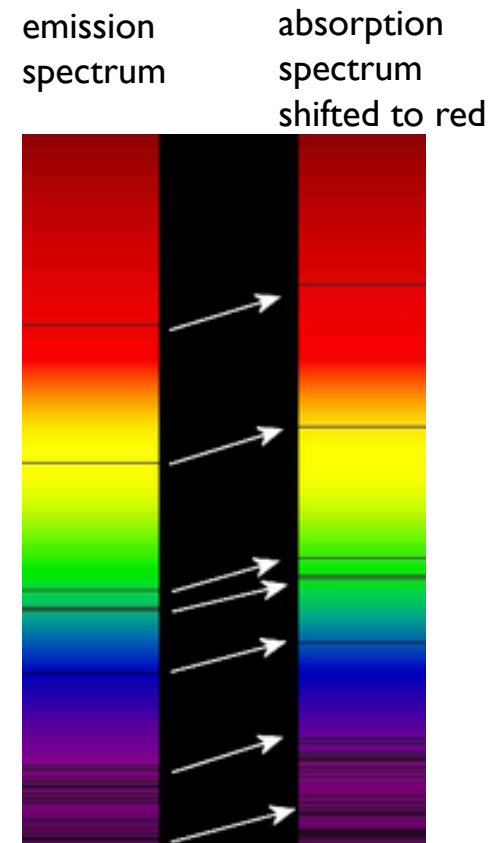
Specifically, we can couple our knowledge of emission spectra as found experimentally on earth, absorption spectra as generated by star-light through spectroscopes and the Doppler Effect to determine how fast stars are moving toward or away from us.

How so?

If you have done as asked and now understand blackbody radiation inside a star, you will understand that radiation frequencies generated only at deeper levels will be more intense than radiation frequencies generated from the surface of the star. This is the origin of the variation in frequency intensity that shows itself as a star's absorption spectra.

So here's the scenario. A star puts out light that has inherently within it an absorption spectra, depending upon which elements are active for a star of that temperature. The light travels to our little rock (earth) and one of our telescopes gathers that light. The light is put through a spectroscope and its absorption spectra is exposed.

What we find, when we do this, is that although the spectral lines from different elements show up as absent, they don't exactly match the lines we have from experimentation here on earth. The spacings are correct, but the actual frequencies have been shifted.



(figure from Wikipedia)

And what is the origin of this shift?

The shift is a Doppler shifted. It is due to the star's motion relative to us.

We have a Doppler relationship that links actual frequencies (we know these because we know the emission spectra for the elements whose lines are missing), observed frequencies (we get this from the absorption spectrum) and relative speed. Using that we can determine how fast the object is moving, relative to us.

If the object is moving away from us, we will observe a shift toward the red end of the spectrum (this is called *red shift*). If the object is moving toward us, we will observe a shift toward the blue end of the spectrum (this is called *blue shift*).

Just so you know, what we observe from all object large distances away are shifts to the red. This is how we've concluded that everything is moving away from us.