Name:____________________ TA:___________________

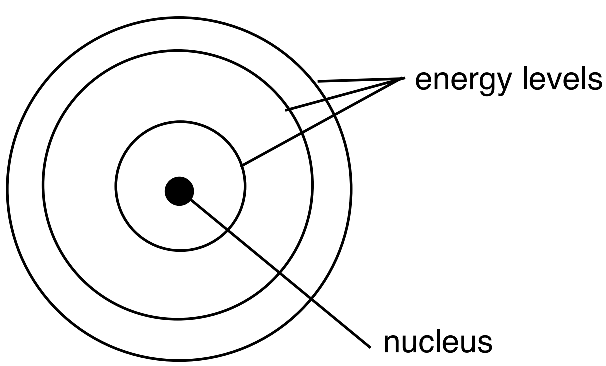
Physics 10263 Lab #2:
<u>Spectroscopy</u>

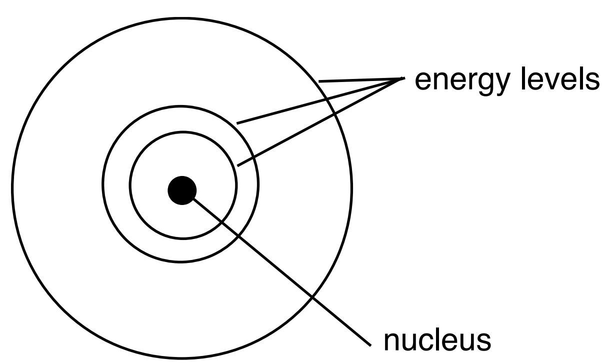
<u>Introduction</u>

In this lab, we'll learn about how emission and absorption works, and we'll learn how to identify elements by observing their emission line "fingerprints".

First, let's talk about atoms and energy. Put very simply, an atom consists of a collection of protons and neutrons at the center called the nucleus. Somewhere in a sort of orbit around the nucleus you'll find electrons. An amazing fact about electrons in atoms is that they are only allowed to have specific values of energy. These precise values depend upon the nature of the nucleus, so the values of energy that an electron is "allowed" to have vary from element to element (i.e. the values for Hydrogen are different than the values for, say, Neon).

Another way of describing this is to say that electrons are only allowed to have certain "orbits" (or "energy levels") about the nucleus, and the precise orbital distances (hence, energies) vary from element to element. A simple model of an atom looks something like one of these two diagrams:





Note that each of these diagrams represents a different element. Electrons in the element on the left have a different set of "allowed energies" compared to electrons in the element on the right.

Important Rule: Electrons are only allowed to exist at certain energy levels within an atom. They aren't allowed to exist "between" levels. Thus, in the atom on the left, electrons might be allowed to have energies of 4, 8 and 11. In the atom on the right, they might be allowed to have energies of 3, 4 and 12.

Why is this important? Because atoms interact with light. By watching the interaction of atoms and light, we can determine the identity of the atom. This careful study of the light emitted and absorbed by atoms is called <u>Spectroscopy</u>. It involves splitting light up into its component wavelengths, or colors (the spectrum) and studying what is happening to the light at each wavelength.

Suppose an atom's electron is in an "excited state", which simply means its electron is at some energy level besides the lowest level. At any given time, the electron may spontaneously decide to jump down one or more levels. Note well that the electron cannot just move at random, it has to move from one energy level to another. For example, in the atom on the right, suppose the electron is in energy level 12. It can move down to either level 3 or level 4. By doing so, the electron must emit the amount of energy it loses.

Thus, if an electron drops from level 12 to level 4, it must give off (12 - 4 =) 8 energy in the form of light. If it drops from level 12 to level 3, it must give off 9 energy in the form of light. From the text and from lecture, we know that the energy of light is inversely proportional to its wavelength.

So, let's summarize how atoms emit light:

1) Electrons are excited to high energy levels (usually by heating up the atoms).

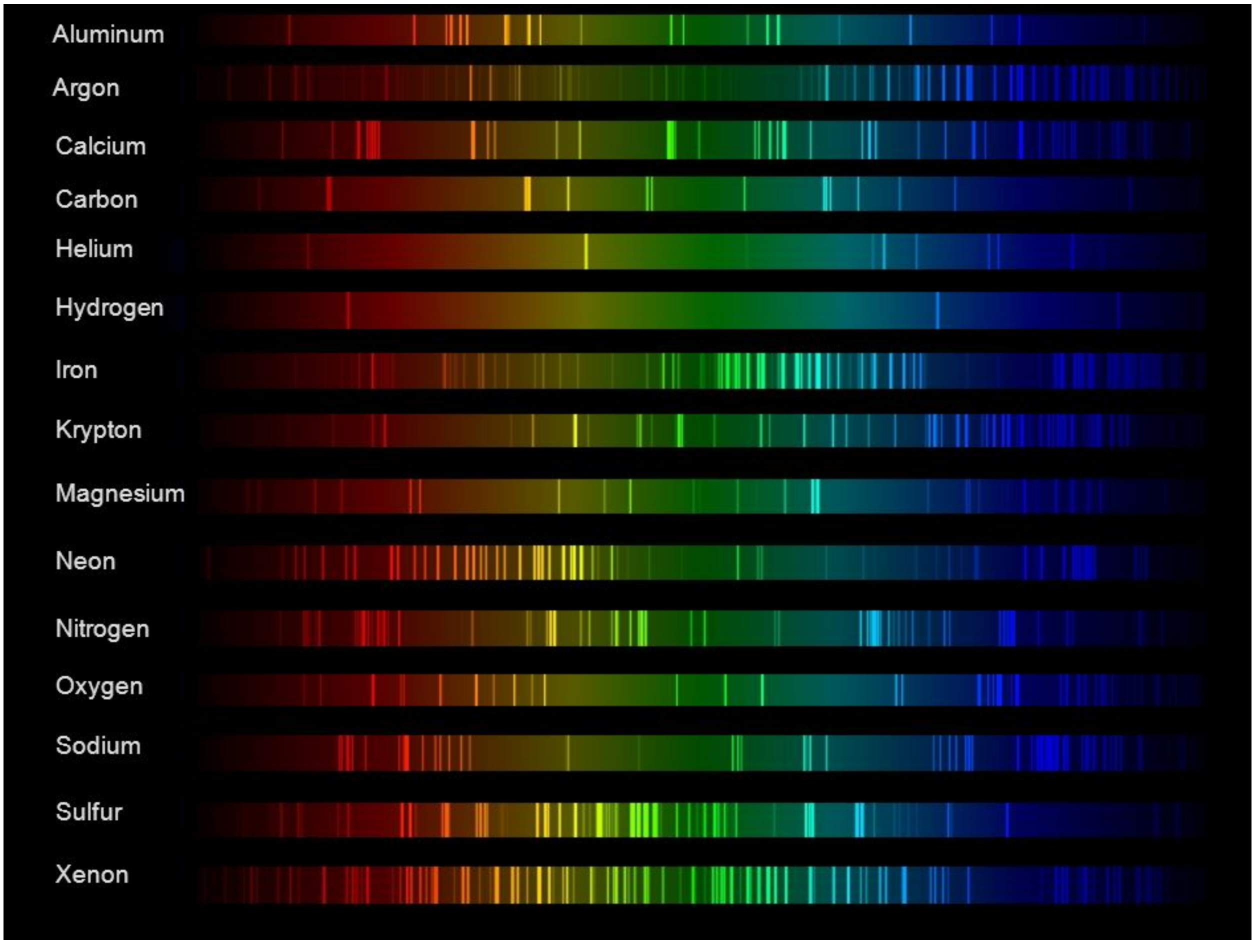
2) These electrons eventually drop down to lower energy levels, giving off light with an energy equal to the energy difference between the higher and the lower level.

3) Since every atom has a unique set of energy levels, the electrons in these atoms will give off light with a unique set of energies (corresponding to energy differences between the various levels).

4) Since energy is inversely proportional to wavelength, each atom will give off light with a certain "pattern" of energies or wavelengths.

So by breaking up light into its various wavelengths, we can identify the atom that is emitting the light, because each atom will have a unique pattern (some examples are shown below).

In some cases, emission lines can be broadened, not sharp and exactly at one particular energy like you might expect. This can be due to a variety of factors such as temperature, pressure and abundance.



Note that these emission spectrum are only valid under certain conditions of temperature and pressure. If the temperature were to increase, for example, the electrons might have enough energy to exist at higher energy levels and so make larger energy (shorter, bluer wavelength) jumps. So some of the lines on the red end of the spectrum might get weaker, and some lines on the blue end might get stronger (and some new lines we don't see now may appear). Similar changes may occur if temperature were to decrease or pressure were to change.

So when trying to identify a particular gas, we usually just look for the strongest 2-5 features in the spectrum. For example, let's suppose you look at the spectrum of some unknown gas, and it looks like this:

What would you identify this as? It's not an absolute perfect fit to any of the reference spectra on the previous page, but there is a match there. This is a great example to discuss with your classmates and your TA during the introductory part of the lab.

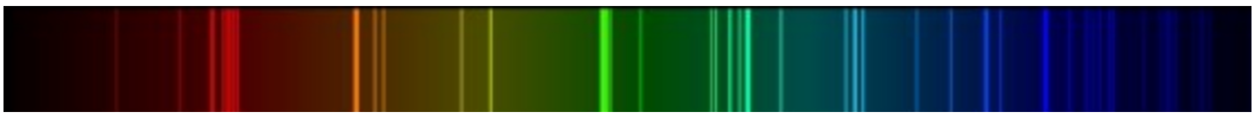
Once you have identified it, (a) state what element you think it is and (b) on the image above, mark the bright lines (with a little notch or check mark above the image) that match most closely with the reference spectrum of your element.

Unknown element #1: ______________________________

If you have printed this and can't see the colors very well in the reference spectra or the sample spectra we are trying to match, then you may need to refer to the pdf on your screen so that the colors are clearer and easier to see. If you are working on your own paper rather than printing the lab, you will need to sketch the sample spectra to mark them.

Now try to identify some other elements from their spectra. If you think the spectral "fingerprint" you are seeing is not on the reference image, just say "unknown," but be sure to discuss each example with others in the group or your lab instructor to be sure!

For each example below, (a) state what element you think it is and (b) on the image, mark the bright lines (with a notch or check mark above the image) that match most closely with the reference spectrum of your element.



Unknown element #2: ______________________________

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Unknown element #3: ______________________________

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Unknown element #4: ______________________________

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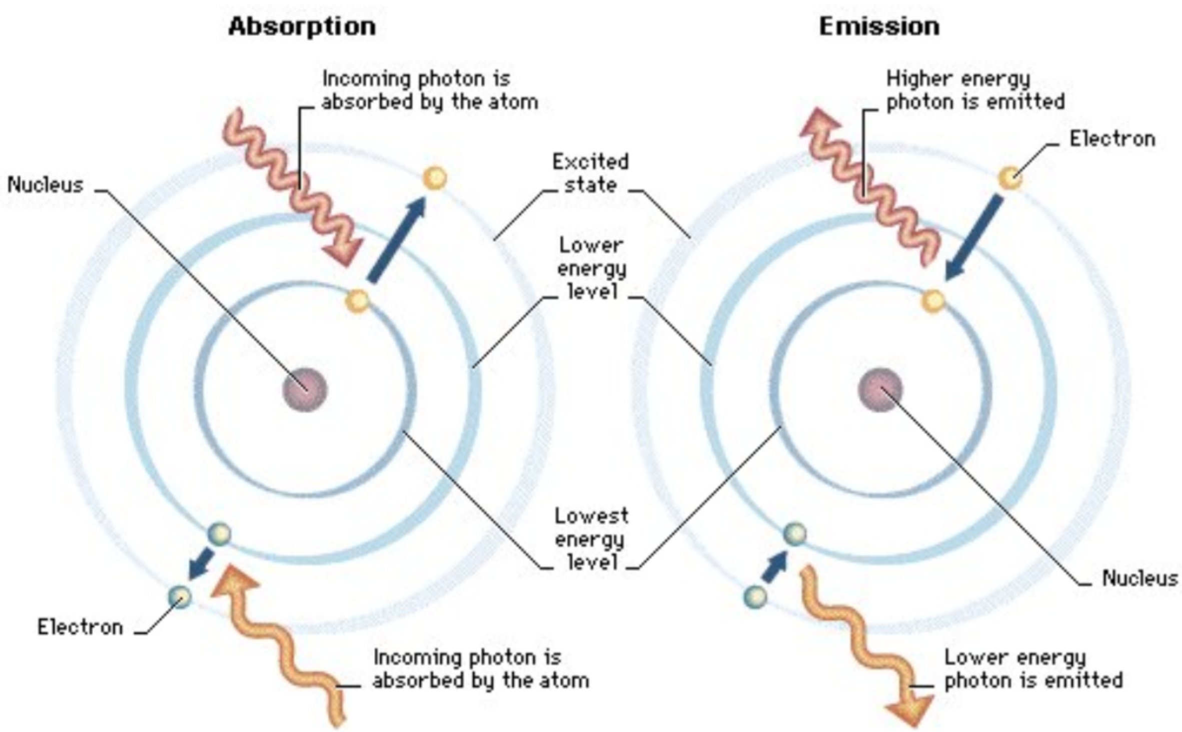
Unknown element #5: ______________________________

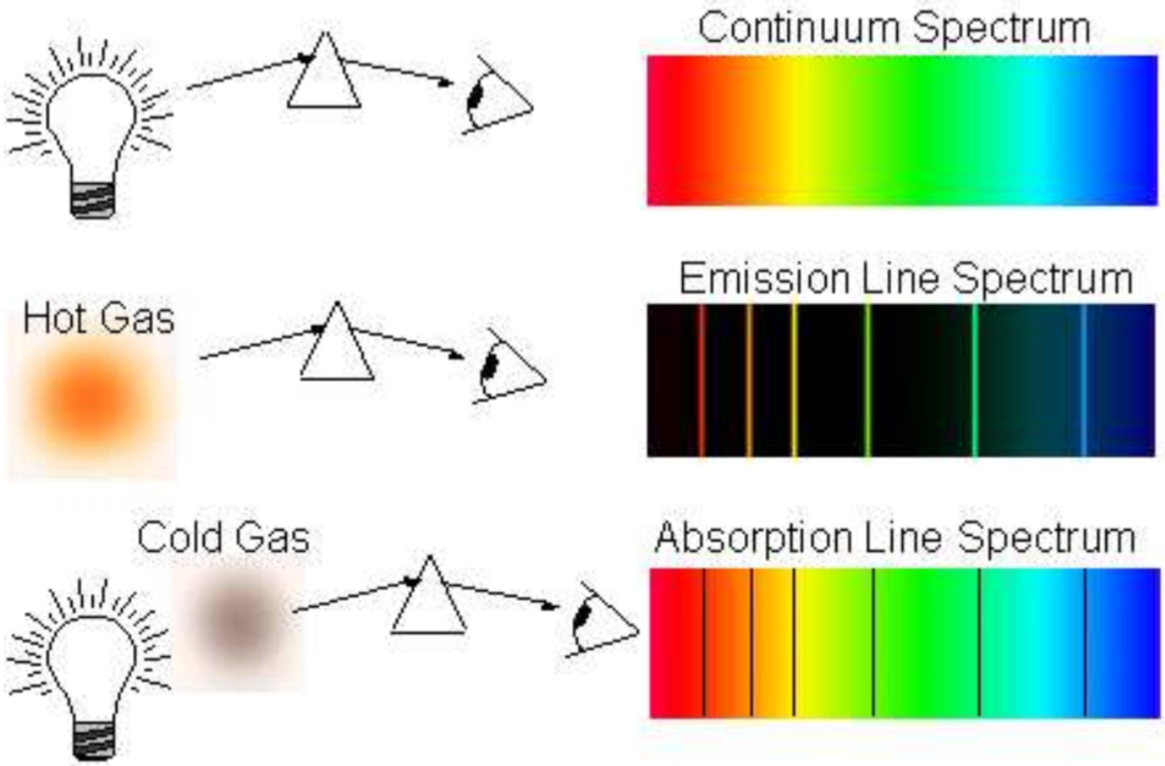
Astronomers use atomic spectra to study the compositions of stars that are sometimes billions of light years away. When you stop and think about it, it is pretty amazing that we can determine the composition of an object so far away without ever touching it or interacting with it. Without the science of spectroscopy, we wouldn't know much of anything about what other planets, stars or galaxies are made of.

Now that we have seen examples of how to identify elements by their emission lines, let's talk about absorption lines.



In the diagram above, you can see what happens at the atomic level when atoms emit light, resulting in the type of spectra we have already seen. The electrons get excited somehow (by absorbing light or perhaps through collisions with other atoms in a hot gas) and when they drop down to lower energy levels, they emit the patterns of light we have seen.

Conversely, on the left, when electrons are in lower energy levels, they may absorb certain wavelengths (or energies) of light, jumping up to higher energy levels. Again, only specific wavelengths can be absorbed, so there is an identifiable "fingerprint" in both emission and absorption.



The diagram above summarizes three basic types of spectra and under what conditions we see them. First is a simple continuous rainbow spectrum without any absorbing or emitting clouds of atoms in between us and the source.

In the second image, we are looking at a gas cloud (or gas glowing in an emission tube) that has been heated, and its atoms are emitting light in the kinds of patterns we have seen.

In the third image, a gas cloud is in the way, between us and some source of continuous light. In this case, the atoms in the clouds will absorb some of the light, and the electrons in these atoms will be excited. Note that the absorbed wavelengths (or energies) precisely match the pattern of the emission line spectrum, assuming it is the same element.

Below is a spectrum of starlight. The light from this star’s surface has passed through relatively cooler gas above the surface, and that gas has left a spectral fingerprint in the absorption pattern. Can you identify what gas this is based on the location of the absorption lines in comparison with the emission patterns in our reference spectra? Just focus on the 4-5 strongest features.

What is this unknown element #6? ______________________________

As we mentioned earlier, spectral fingerprints can change based upon conditions changing like temperature or pressure. The absorption spectrum below is for a different star with the same composition as the star with unknown element #6:

Notice that the spectrum looks almost completely different. The absorption lines of the unknown element are almost completely gone, and yet that element is still present in the star! It’s just that, in this case, the star is not hot enough to excite the electrons in that element, so they never participate in the absorption or emission process.

So the science of spectroscopy is a little more complicated than simple pattern matching. We have to take into account that the pattern displayed by an element depends on the properties of the star, nebula or galaxy that we are studying, such a temperature, pressure, density, magnetic fields and motion. Not to worry, Astronomers have gotten really good at this over the decades since we started analyzing spectra. You’ll learn throughout the semester just what we can learn from light!

<u>Essay</u>: To conclude this lab, please write a short essay on your own paper that explains (1) how a spectral fingerprint for an element is formed, (2) why each element has a unique fingerprint and (3) why the element’s fingerprint can change if the temperature changes.