Activity 9: Atoms and Spectra

Materials:

- Ruler
- Incandescent white light bulb (at front of classroom)
- Spectroscope or diffraction grating film
- 5 numbered gas lamps (distributed around classroom)
- Hydrogen gas lamp (at front of classroom)

Starting with today's Activity, we will be shifting our focus in this course from studying and interpreting what we see in the sky to addressing another question: "How do we know what we know about things that are so far away?" As discussed earlier in this course, even the nearest star system, Alpha Centauri, is too far away for us to visit it. So if we can't go there, how can we learn about it?

In a word, the answer is: *light*. Nearly everything we know about objects in the Universe comes to us via the light that reaches us from those objects. Astronomers are like Sherlock Holmes: they can deduce a great deal from just a few small clues. Today, we will start to unravel this detective work as we take a look at how and why information is encoded in light, and how we can learn something from it.

Part 1: Electromagnetic Spectrum Review and Spectroscopes

Frequency and Wavelength:

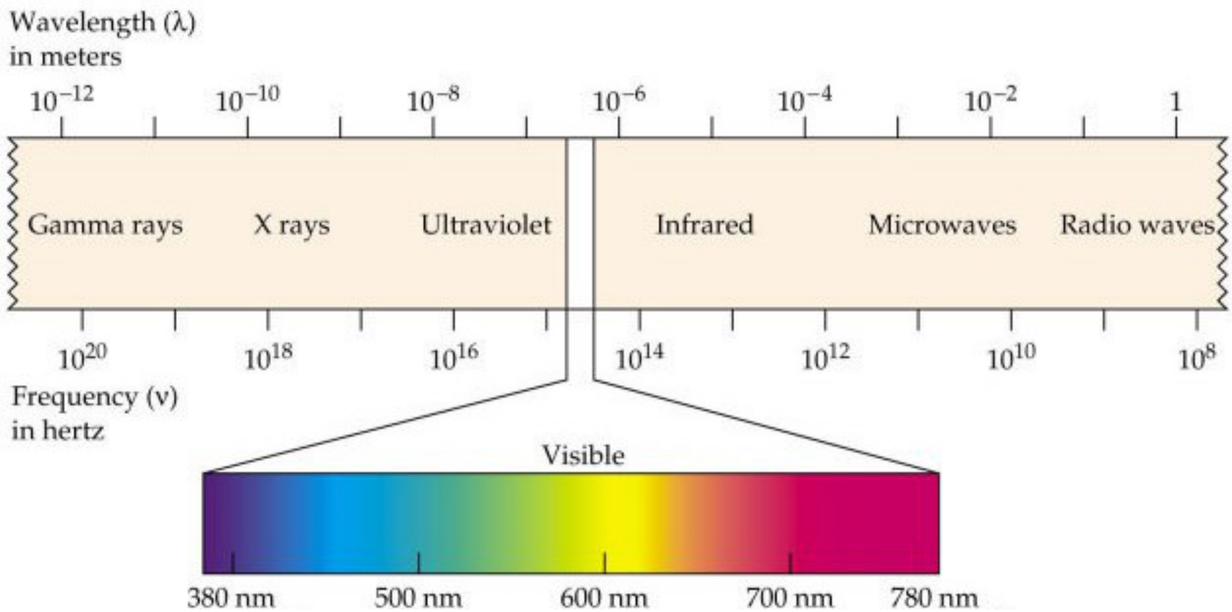
a) As the *wavelength* of a wave *increases*, what happens to the wave's *frequency*?

b) As the wavelength of a wave *decreases*, what happens to the wave's frequency?

c) Based on your answers to (a) and (b), what kind of relationship exists between frequency and wavelength (direct or inverse)?

d) What kind of relationship (direct or inverse) exists between a wave's frequency and a wave's energy?

Use the electromagnetic spectrum below to answer the questions that follow. A color version of the visible spectrum can be found here: tinyurl.com/276w523d



What is the most energetic form of electromagnetic radiation?

What is the least energetic form of electromagnetic radiation?

What is the most energetic color of visible light?

What is the least energetic color of visible light?

What is the *smallest wavelength* (approximately, in nanometers) that the human eye can see?

What is the *largest wavelength* (approximately, in nanometers) that the human eye can see?

STEP 1: Let's take look at a regular incandescent light bulb. This bulb emits what we call *white light*.

Using a spectroscope or diffraction grating film, view the light bulb by looking through the spectroscope or film.

How to use a spectroscope:

> Notice one end of the spectroscope has an eyehole, and the other end of the spectroscope has a vertical slit that acts as a viewing window. That viewing window (the vertical slit) should face a light source at the same time as you look into the eyehole on the opposite end.
>
> Point the spectroscope's viewing window at the light bulb and look into the spectroscope. You may have to move your head left and right to see everything inside the spectroscope, while keeping your line of sight aligned with the eyehole.
>
> Inside the spectroscope, you should see the unaltered light of your target light bulb on one side, shining through the vertical slit.
>
> On the other side, you will see the spectrum of that light. The spectroscope has taken the white light and split it into its constituent components.

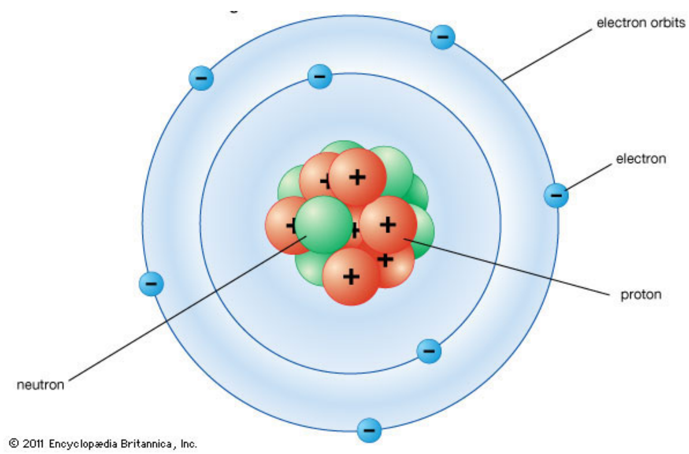
What colors does *white light* consist of? To answer this, look through the spectroscope or film and see what colors the white light of the bulb has been split into. Write them below, in the proper order, as you see them from left to right. Don't worry about various shades, just stick with basic color names (red, blue, yellow, etc).

Does your list of colors agree with the mnemonic ROYGBV (perhaps in reverse order)? If not, try again or ask your instructor for assistance.

STEP 2: Look inside the spectroscope and notice the numbers that are listed along the bottom (if any), beneath the spectrum of light that you see. Looking back at the electromagnetic spectrum above, what do you think these numbers represent?

Part 2: Atoms and Emission Spectra

STEP 1: You are probably familiar with drawings of atoms that look like this:



Notice that the electrons have orbits around the nucleus. Electrons cannot just orbit anywhere they like. They must orbit in one of several discrete orbits and they cannot be in-between orbits. This is the basic premise of **quantum physics**: electrons can't have just any energy, because energy comes in discrete chunks (a single chunk of energy is called a *quantum* of energy).

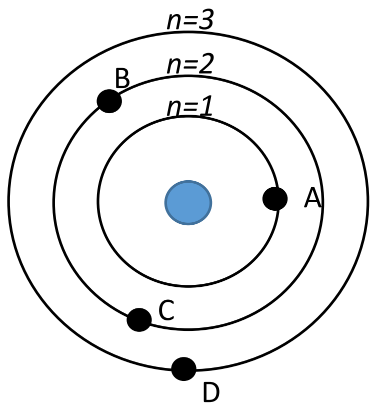
The following bullet points about electron orbits are fundamental in understanding what we'll be doing for the rest of this Activity:

- All of the electrons in a given orbit have the *same* amount of energy.
- The *largest* orbit is the one in which the electrons have the *most* energy.
- The *smallest* orbit is the one in which the electrons have the *least* energy.
- These orbits are given the name *energy levels* because the energy of an electron is determined entirely by which energy level it is in.

Conventionally, the energy levels are numbered, with $n=1$ being the lowest energy level (smallest orbit). $n=1$ is often called the **ground state** for electrons.

There is no limit to the number of energy levels an atom may have. In this Activity we will mostly be looking at diagrams that have between 3 and 6 energy levels. However, any atom can have any number of energy levels.

SQ1:



a) In the drawing of the atom on the left, which electron (A through D) has the most energy? Which has the least?

Most energy:

Least energy:

b) In the drawing, which electrons have the SAME energy as each other? How can you tell?

c) Suppose an additional electron were added to the above atom. If it winds up on the n=3 energy level, it will have the same energy as which electron that's already shown?

Electrons can jump from one energy level to another. For instance, electron D in the above drawing can move from the n=3 energy level to either the n=2 or n=1 energy levels.

Which energy level(s) of those depicted above would electron C be able to jump to?

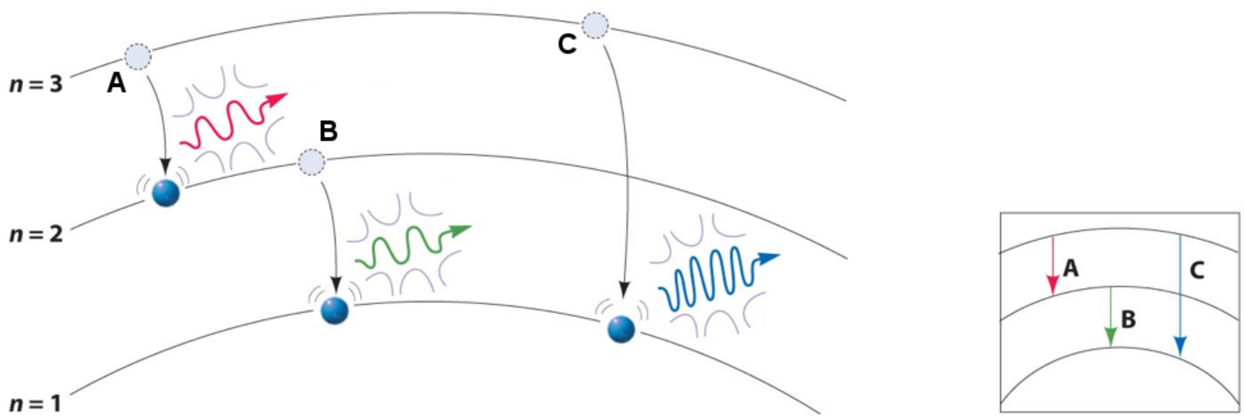
> Aside: Since electrons can be *in* orbits but not *between* orbits, when an electron changes energy levels, it jumps *directly* from one orbit to another, without traversing any of the energies in-between orbits. This hard-to-fathom phenomenon has been dubbed the *quantum leap*.

If an electron drops down in energy, what happens to that energy? Let's discuss that now.

Conservation of energy is the law of physics that says that energy cannot be created or destroyed.

Therefore, when an electron drops down in energy, that energy must *go* somewhere because it *can't* be destroyed. So where does it go? The answer is: it is emitted as electromagnetic radiation!

STEP 2: Consider the diagram below. It shows three electrons making a transition from a higher energy level to a lower one (as indicated by the arrows). The amount of energy each electron loses is the amount of energy that each one releases as an electromagnetic wave. Important note: This diagram is *not* to scale because, in a real atom, the energy levels are *not* evenly spaced.



SQ2:

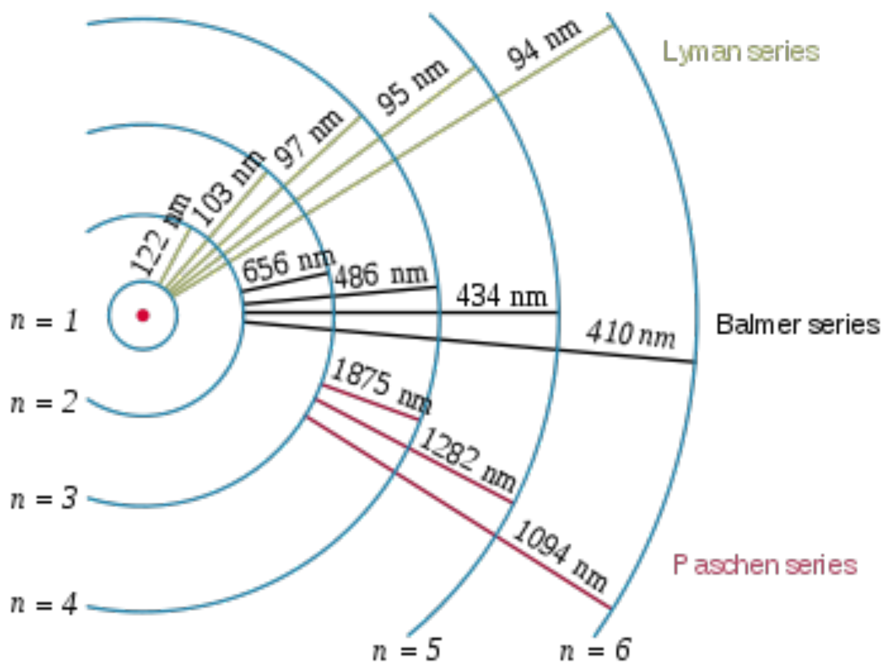
a) Which electron releases more energy, A or C? How can you tell from their change in energy levels? (Remember, energy levels are *not actually* evenly spaced, but you can still draw conclusions such as: electrons at n=3 will have more energy than electrons at n=2, and so on.)

b) What does this tell you about the frequency of the radiation that's emitted by A as it drops versus the frequency of the radiation that's emitted by C as it drops?

c) What do you think is the relationship between the energies released by A, B, and C? Can you write down this relationship as a simple equation?

STEP 3: Below is a diagram of some of the energy levels of a **hydrogen** atom. In the diagram, numerous possible electrons are shown dropping from their starting energy levels down to lower energy levels.

Energy Levels of a Hydrogen Atom



The wavelength of the light emitted by each transitioning electron is written in the diagram. The wavelengths are given in nanometers (nm).

Are the energy levels of hydrogen equally spaced in terms of energy? Let's use the diagram to find out.

a) If an electron drops from n=2 to n=1, look at the diagram to see what wavelength of radiation will be emitted. What is that wavelength?

Radiation with this wavelength has energy equal to the difference between n=2 and n=1.

b) And the energy difference between n=3 and n=2 is equal to the energy of radiation with what wavelength?

c) Does the radiation emitted in (a) have the same energy as the radiation emitted in (b)? If not, which has more energy?

d) Which change in energy levels is a bigger change in energy, between n=1 and n=2, or between n=2 and n=3?

The electron transitions in hydrogen can be grouped into categories, called “series”. Look back at the diagram to answer the following questions:

a) The *Lyman series* consists of all of the transitions in which an electron drops down to which energy level?

b) The *Balmer series* consists of all of the transitions in which an electron drops down to which energy level?

c) The *Paschen series* consists of all of the transitions in which an electron drops down to which energy level?

Looking at all of the different wavelengths of light being emitted by a hydrogen atom, would any of these be visible to the human eye? If so, is there a particular series (Lyman, Balmer, Paschen) of transitions that seems like it would be especially well-suited to being seen by the human eye? (Hint: Back on p. 2 you already figured out the smallest and largest wavelengths the human eye can see.)

STEP 4: Now let’s consider the colors of light that would be produced by the series that you wrote down in the previous question.

If you need a color image of the visible spectrum and its wavelengths, open this URL:

tinyurl.com/276w523d

Compare the wavelengths emitted by the series you named earlier with the wavelengths of different colors of visible light as seen at the URL you just opened. Which colors would you expect to see with your own eyes as a result of the electrons dropping energy levels within hydrogen atoms? Just write down the general color name (red, green, etc.) you’d expect to see.

Colors you’d expect to see:

STEP 5: At the front of the room is a lamp with a tube of hydrogen gas in it. Ask your instructor to turn on the lamp. The hydrogen inside the lamp will heat up and glow. What color does the lamp appear to be when you look at it with just your naked eye?

Remember the light bulb: what looked like white light to our eyes was actually light that consisted of all colors. Similarly, the light you see here from the hydrogen lamp may actually be a combination of light of various colors.

STEP 6: Let's see what the colors are that make up the light that you are seeing from the lamp. Look at the lamp using your spectroscope.

What colors do you see through the spectroscope? (If you don't have a hydrogen lamp, ask your instructor which colors you would see when viewing a hydrogen lamp through a spectroscope.)

This is called the **emission spectrum** of hydrogen. Did it include the colors you'd predicted a hydrogen atom would emit? Were there differences between your observation and your prediction?

Because each atom has a differently structured nucleus, it also has a different arrangement of energy levels around the nucleus. Therefore, the arrangement of energy levels of each element on the periodic table is *unique*, and hence so is the emission spectrum of each element!

The emission spectrum of an atom is like a fingerprint that tells us what element the atom is.

STEP 7: To see the emission spectra for various elements, open this URL:

astrodave.name

(If you do not have any gas lamps, you can skip the following procedure and continue on to Step 8).

PROCEDURE: There are **five** numbered tubes of gas in the classroom. Your goal is to identify what gas is in each of the tubes. You can do this by pointing your spectroscope at each gas tube and comparing what you see with the spectra at the URL you opened above. The five gas tubes are distributed around the room for everyone to use. So that we all use the same numbering, on top of the black portion of each lamp is a piece of tape with a number 1-5.

Important: Do not remove the gas tubes from the lamp. Also, do not leave the lamp on for more than 30 seconds at a time because they can burn out. **Remember: turn off each lamp before you walk away from it**.

Hint: Be very careful with Krypton and Mercury. Their spectra look very similar, so look carefully when you compare them with what you see in the spectroscope.

SQ3: Below record your findings regarding which kind of gas you've concluded is in each of the five tubes.

1: 4:

2: 5:

3:

STEP 8: What would happen if we put multiple gases in the same tube? For instance, if a tube contained both mercury and argon, what would its emission spectrum look like? Describe what colors you'd expect to see in the spectrum.

Part 3: Continuous and Absorption Spectra [1]

STEP 1: So far we've been considering electrons that drop down to lower energy levels. But electrons can climb to higher energy levels, too.

If an electron climbs to a higher energy level, it *gained energy*. Conservation of energy then tells us that the added energy must've *come from somewhere* because energy can't be created out of nowhere.

Describe how this process might work. What is the incoming energy? What happens to it, and what happens to the electron? *Hint:* It is the same process that we saw earlier when an electron drops to a lower energy level, except everything happens in reverse!

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[1] Images in this Part are from *Lecture-Tutorials for Introductory Astronomy* (3rd Edition) by Prather, et al.

STEP 2(a): Look back at the "Energy Levels of Hydrogen" diagram. The wavelengths indicated on the diagram are the wavelengths of light that a certain electron would emit when it drops energy. To review, if an electron drops from n=3 to n=2 in a hydrogen atom, it releases radiation with a wavelength of what?

And this is because radiation with that certain wavelength (and therefore a certain frequency, and therefore a certain energy) exactly makes up the difference in energies between the two energy levels.

STEP 2(b): If an electron were to *jump up* from n=2 to n=3, what wavelength of light do you expect it would need to absorb to make up that energy difference?

Were your answers to Steps 2(a) and 2(b) the same? Explain why in your own words.

SQ4:

a) For the hydrogen atom, what wavelength of radiation would you expect an electron to absorb in order to *climb up* from n=2 to n=4? Look back at the Energy Levels of Hydrogen diagram.

b) What about from n=1 to n=4?

c) Which of these two wavelengths has a higher frequency? Therefore, which has a higher energy? Why does this make sense with respect to how much the electrons changed their energy levels?

SQ5:

(a) If all of the electrons in an atom are in the lowest energy level, will that atom be able to *emit* much radiation? Why or why not?

(b) If all of the electrons in an atom are in the lowest energy level, will that atom be able to *absorb* much radiation? Why or why not?

(c) If most of the electrons in an atom are in high energy levels, will that atom be able to *emit* much radiation? Why or why not?

(d) If most of the electrons in an atom are in high energy levels, will that atom be able to *absorb* much radiation? Why or why not?

STEP 3: Recall in Part 1 when we looked at the light bulb through the spectroscope. Did you see an emission spectrum that included lines of only a few specific colors, like in an emission spectrum? If not, then what did you see?

When we see this type of spectrum, we call it a **continuous spectrum** because it has no gaps and it encompasses all wavelengths. A continuous spectrum includes *"all of the colors of the rainbow"*.

Suppose light of all wavelengths (i.e., a continuous spectrum of light, or "all of the colors of the rainbow") encounters a gas. If all (or most) of the atoms in that gas contain electrons that are in high energy levels, will the light passing through the gas be absorbed much? Why or why not? Hint: Look back at your answers in SQ5.

If radiation of a certain wavelength isn't absorbed by an object, then the radiation passes right through it. We then say that the object is ***transparent to*** radiation of that wavelength.

A window is an everyday example of this. Windows don't absorb radiation with visible light wavelengths, therefore visible light passes through and this is why we can see what's on the other side of the window.

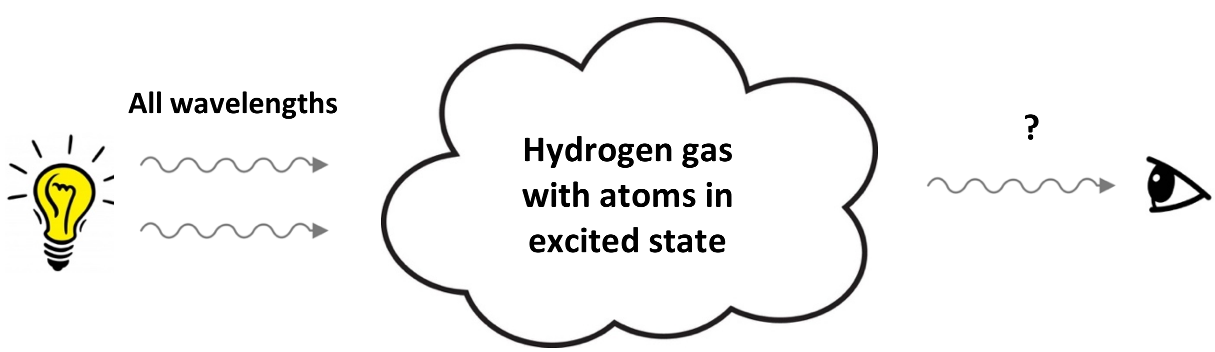
Therefore, a typical glass window is transparent to which part of the spectrum?

The opposite of transparent is **opaque**. An object is *opaque to* a certain wavelength of radiation if it absorbs radiation of that wavelength, hence not allowing it to pass through.

Windows absorb infrared radiation. And windows are often coated with a UV protective coating that also absorbs UV.

For which types of radiation is a typical glass window opaque?

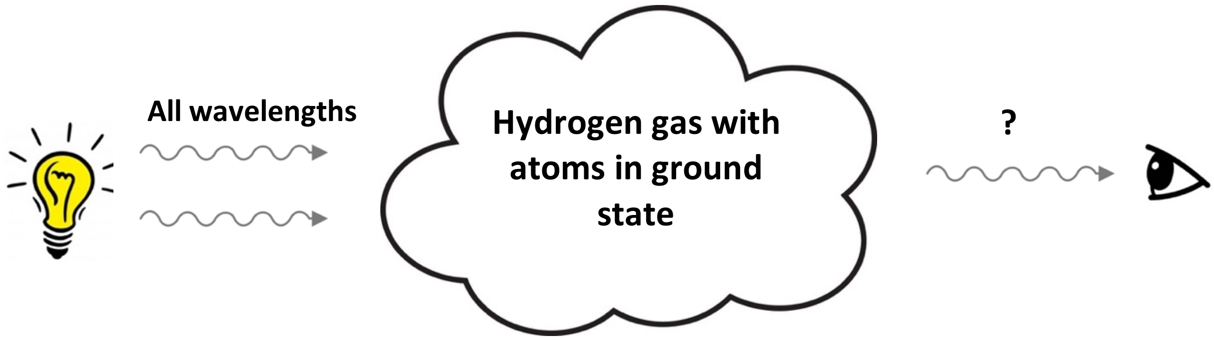
STEP 4: Consider the situation below. There is a light source on the left that emits a continuous spectrum of visible light (like a light bulb). The light (which includes all colors together, AKA "white light") encounters a gas (such as the air in this room, the Earth's atmosphere, or an interstellar cloud).



If the gas is made of hydrogen atoms in which most of the electrons are *not* in their ground state (we call this the "**excited state**"), what colors will the observer on the right see?

Colors observer sees:

Now suppose the gas is made of hydrogen atoms that are mostly in the ground state (i.e., all of the electrons are in their lowest energy levels), as in the diagram below. Now what colors will the observer on the right see?



Colors observer sees:

When all wavelengths of light pass through an object *except* for a few that were absorbed (as with the cloud above), then what we'll see is called an **absorption spectrum**.

If, in the previous example, the observer on the right were using a spectroscope to view the light coming to him/her after passing through the cloud, what would the spectrum look like to that observer?

STEP 5: Ask your instructor for the color handout of the absorption spectrum of a hydrogen gas. Describe its appearance here.

SQ6: The dark bands exist where the colors were absorbed by the hydrogen gas and thus never made it to the observer. Based on what we've learned about energy levels in atoms, why do you think it is *those* colors in particular that are missing?

SQ7:

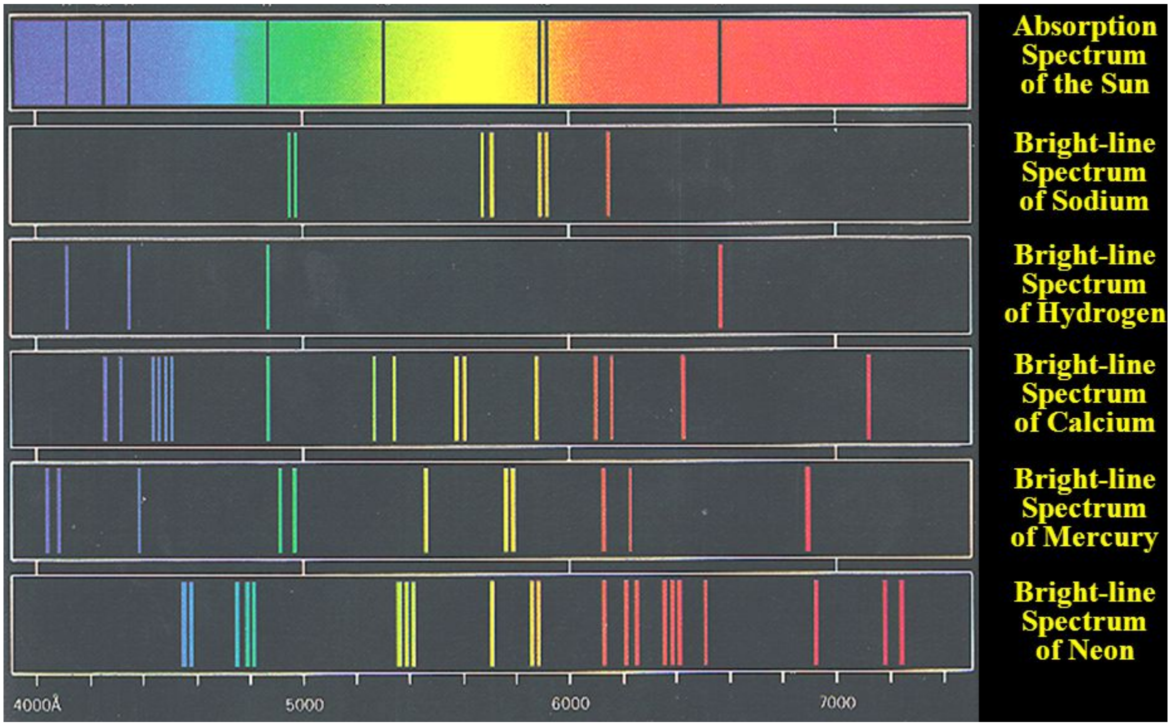
Stars emit a *continuous spectrum* of radiation because they are hot and dense. But stars also have *atmospheres* around them. An atmosphere is a low-density layer of gases that is cool enough that *most of its atoms will be in the ground state*.

a) If you could find a star that had no atmosphere around it, which type of spectrum would you expect to see from the star (continuous, emission, absorption)?

b) When you look at the Sun or another star that has an atmosphere, what kind of spectrum would you expect to see (continuous, emission, absorption)? Explain your reasoning. (Hint: Compare this situation to one of the light-bulb-and-cloud examples above. How are they similar?)

c) Would you expect *Earth's* atmosphere to play a role in the spectra that we see when using telescopes on Earth's surface? Why or why not?

STEP 6: The top row in the image below is the absorption spectrum of the Sun. Looking at the absorption lines, and comparing them to the emission lines for various elements, can you identify a few gases that are present in the Sun's atmosphere? Hint: It might be helpful to lay a ruler vertically on the diagram to help you decide whether things line up or not.



Using the above spectra, can you identify some of the gases that are present in the Sun's atmosphere? Write them below. Then check your answers with your instructor.

The above absorption spectrum for the Sun is highly simplified. There are actually *dozens* of absorption lines in the Sun's spectrum. To see a more detailed image of the Sun's absorption spectrum, open this URL: tinyurl.com/3tc552pn