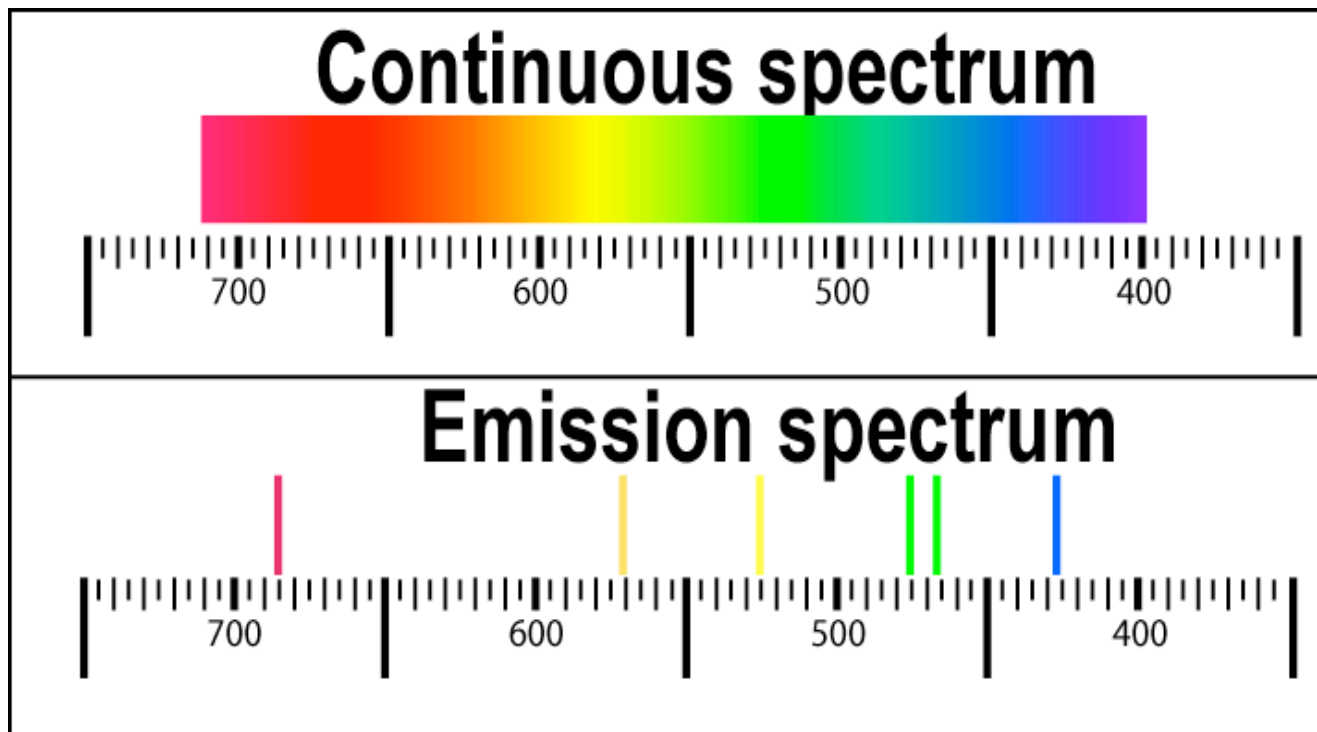


Introduction to spectroscopy

How do we know what the stars or the Sun are made of?

The light of celestial objects contains much information hidden in its detailed color structure. In this lab we will separate the light from some sources into constituent colors and use spectroscopy to find out the chemical constitution of known and unknown gases. The same procedure is used for starlight, telling us what its source is composed of. The baseline is a laboratory experiment with known materials, and later we can compare the unknown to what we already know.



Hot, glowing bodies like a light bulb, or the Sun, glow in all the colors of the spectrum. All these colors together appear as white light. When such white light hits a prism, or a raindrop, or a diffraction grating, colors get separated according to their wavelength. Red, with its wavelength of 600 nm to 700 nm, is deflected least and ends up on one edge of the spectrum – or the rainbow when sunlight hits a raindrop after a storm. Blue, wavelength around 400 nm , is the other end of the visible spectrum. An infinite number of elementary colors are located between these two edges, each corresponding to its own wavelength.

An incandescent light bulb radiates a continuous spectrum. All colors are present in this “thermal glow”, and it is impossible to tell what the chemical composition of the source is. However, other physical processes produce different spectra. A fluorescent light tube works, crudely speaking, on the principle of lightening. Electrons rush from the negative pole to the positive pole inside, and hit gas atoms in the tube, making them emit light. This sort of light contains only a few colors, and is called “emission spectrum”. When we separate the colors of such light, only a few bright “emission” lines appear, each in its own color (and wavelength). Each sort of an atom will emit light at its own particular set of wavelengths. When we analyze

the emission spectrum of an unknown source, we can compare the colors of its spectral lines to known spectral lines we see in a laboratory, and tell which substance matches.

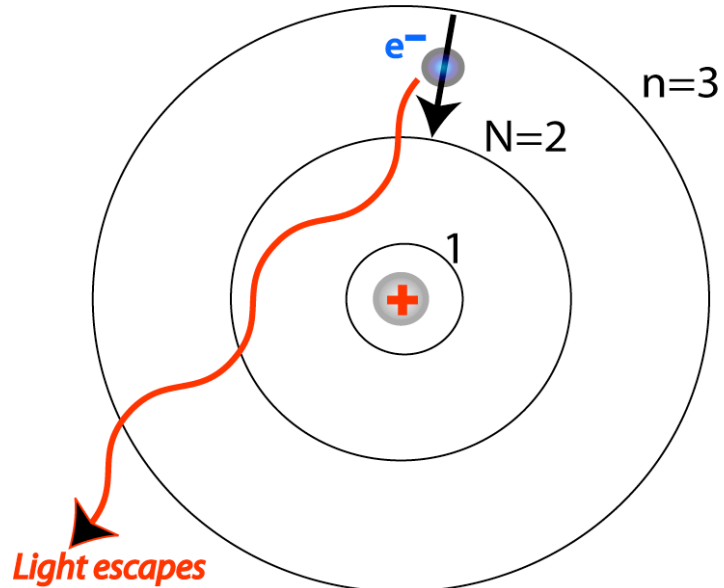
Having read this much, please answer the first three questions, and write your answer in the space provided on your report sheet.

The color of spectral lines is directly related to the structure of the atoms. Electrons can jump from one orbit around the atomic nucleus to another, giving off the difference in energy levels in the form of light. The wavelength (color) of light is inversely proportional to the amount of energy freed up between the old and the new orbit. In the case of hydrogen, there is a simple formula that tells us the wavelength of the spectral lines, called the **Balmer formula**:

$$\lambda = 91.177nm / (1/N^2 - 1/n^2)$$

which tells the wavelength associated with an electron jumping from the N^{th} to the n^{th} orbit. The “Balmer series” spectral, lines which we will measure, called $H\alpha$, $H\beta$, $H\gamma$, ..., correspond to the electron jumping from some level $n=3,4,5,\dots$ down to levels $N=2$. (Of course, hydrogen has another series of spectral lines, the Lyman series, which corresponds to $N=1$, but they are not visible to the human eye.

An electron jumping from $n=3$ to $N=2$.

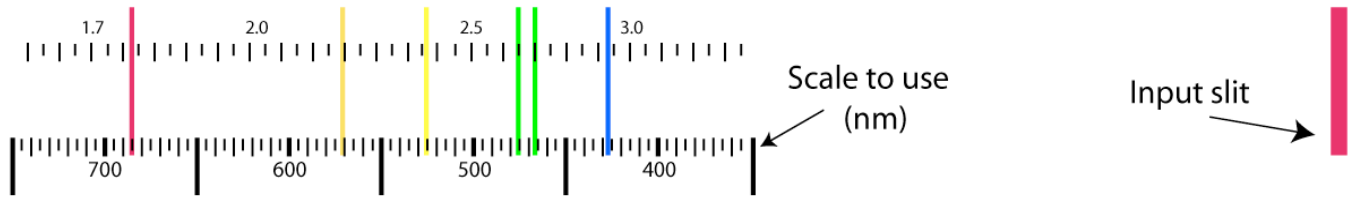


In this laboratory we will measure the wavelengths of spectral lines from a few gases, which are easy to put inside a discharge tube. Other chemical elements would have different spectra, and the details of these spectra also contain information on the temperature, pressure, gravity, speed of motion, and much else inside the source of light. We will apply what we learn here to the study of spectra of stars in a later laboratory exercise.

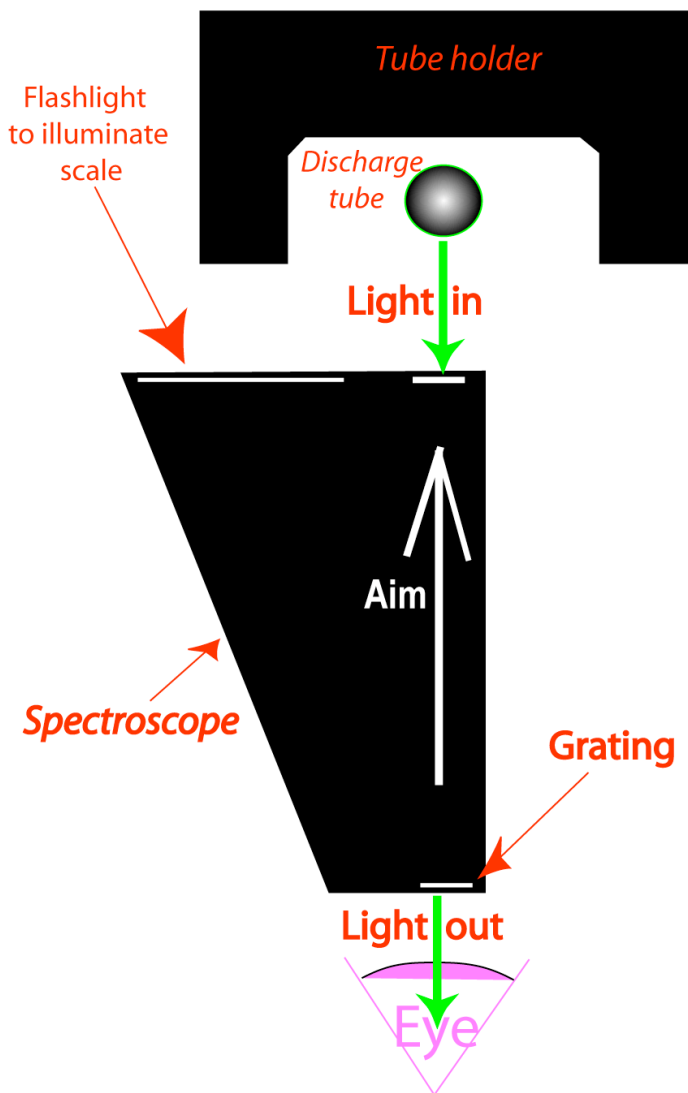
In the spectroscope that we use in this laboratory, a diffraction grating (many parallel black lines drawn very tightly on a little piece of film). It breaks up the light entering through the input slit into colored lines. Each color corresponds to a wavelength, measured in nanometers (nm). Note that $1 \text{ nm} = 10^{-9} \text{ m} = 1/250$ millionth of an inch, a very small unit. The wavelength of visible light is very short indeed.

After reading this, please answer the second set of three questions, and write your answer in the space provided on your report sheet.

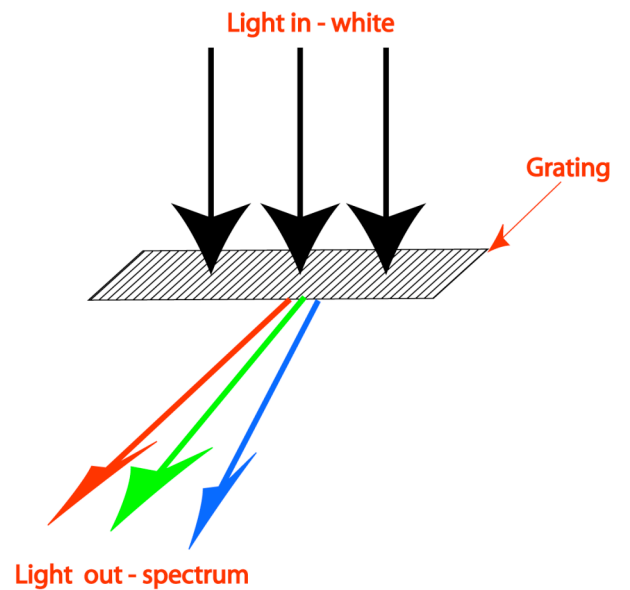
The view in the spectroscope



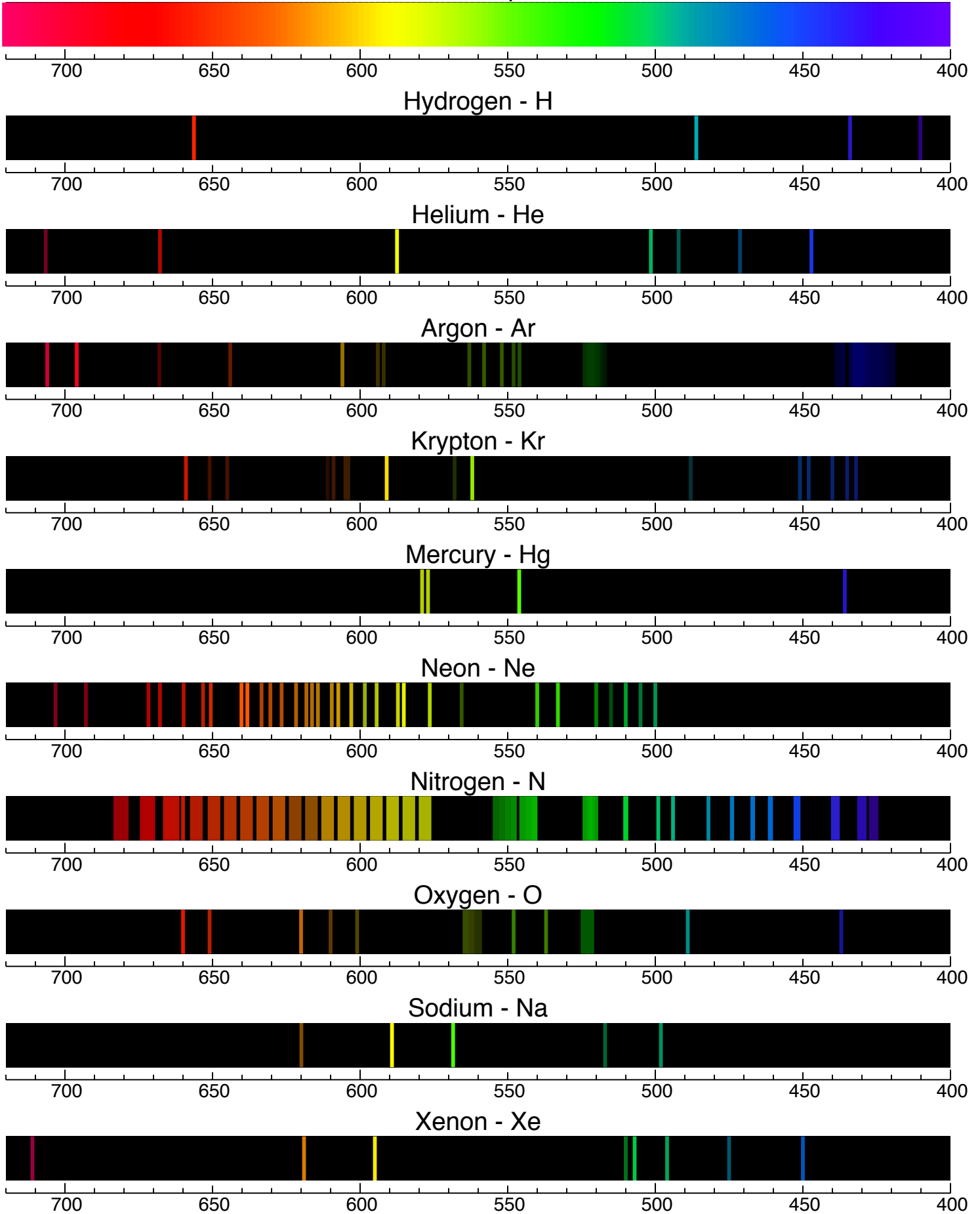
Top view of the setup



A diffraction grating breaks up mixed colors into constituents



Continuous Spectrum



Wavelengths are in nanometers (nm).

Procedure sheet

(1) Sign the roll, form groups of two students, and open the file called “Spectroscopy.pdf” on the computer. Alternatively, you may receive a printed copy from your TA.

(2) Listen to the introduction by your TA.

(3) Read the “Introduction to Spectroscopy” and answer the three questions in the space provided in the report form.

(4) Answer the three questions at the end of the Introduction. These questions ensure you understand what you are reading and get you ready to do the lab.

(5) Examine your spectroscope and identify its parts:

-- Find the opening where light enters. Find the grating and find out how to turn it around. Find out how to aim the spectroscope at a light source.

-- Look into the spectroscope. You'll need to use your glasses or contacts (if any) to see the scales clearly. You will use the wavelength scale (the one that goes from 400 nm to 700 nm) to read off the wavelength of spectral lines.

-- Hold the spectroscope in your hands and aim at a fluorescent light. You'll see the input slit light up on the right (see the picture!). Grab the edges of the grating just under and above the viewing hole to align the grating, so that the spectral lines you see are exactly vertical. If the grating is positioned at the wrong angle, you may not see any spectrum at all.

(6) At this point the ceiling lights will be switched off; plug in the incandescent lamps.

(7) Assemble the setup as shown in the picture. Insert a hydrogen discharge tube into the socket. *Handle the tubes carefully, they are fragile and expensive!* Do not touch the tubes by bare hands: the grease on human hands may cause the tubes blow up, use gloves or a paper towel to touch them. Make sure *the power supply is switched off* to avoid an electric shock! Make sure that the spectroscope is *stable enough* on its holder and that the opening that accepts light aims straight at the middle of the tube. Simply holding the spectroscope in your hands and aiming it at the source will be unlikely to work.

(8) Aim the spectroscope at the discharge tube. First the input slit should light up; make it as bright as possible by aiming carefully. You should now see a few spectral lines (the grating might need a little readjustment at this point). Read off the wavelengths of the three or four bright hydrogen lines (these will be, from red to blue, H α , H β , H γ , and possibly H δ) and record them in your report. If the scale is too dark, you may shine some light into the broad opening on the front left of the spectroscope to illuminate it. Using a color pencil, draw the lines in your report as you see them (*spectrum #1*).

(9) Using the Balmer formula with $N=2$ and $n=3,4,5,6$ calculate the wavelength of each of the lines as predicted by theory and insert your calculated prediction into the table in your report. Calculate the percent error – the difference between theory and experiment. You should expect a 1-2% error. Note that no measurement can ever be absolutely precise, there is always some error.

(10) Replace the hydrogen tube with another three discharge tubes filled with known gases. (*Careful, the tubes are hot!*) Draw a few of their brightest emission lines as you see them (*spectra #3-5*). Use color pencils. With the aid of the laminated spectra, check that the tube contains the correct gas.

(11) Take a “mystery” tube that is not labeled with a name, only with a number. Draw the spectrum of the gas in it, and identify the gas. Use one tube per student, i.e. two per group. (*spectrum #6-7*).

(12) Aim the spectroscope at an incandescent light bulb. Observe that no spectral lines are visible, but you see all colors of a full spectrum instead. Thermal glow produces a continuous spectrum.

(13) Aim your spectroscope at a fluorescent light again. Identify the bright spectral lines with the laminated spectral charts. What gas or gases are there in the light tube? Put your answer in the report.

LAB REPORT: ASTRONOMY 104 -- Spectroscopy

Your name: _____

Section: _____

Date: __/__/200__

Questions answered:

- 1.: _____
- 2.: _____
- 3.: _____
- 4.: _____
- 5.: _____
- 6.: _____

Measured and predicted wavelengths of hydrogen lines

| Name of line | Transit | Predicted λ [nm] | Measured λ [nm] | Color | Error [%] |
|--------------|-----------------|--------------------------|-------------------------|-------|-----------|
| H α | from n=3 to N=2 | | | | |
| H β | from n=4 to N=2 | | | | |
| | | | | | |
| | | | | | |

Drawings of spectral lines and identification of gases

