Experiment 10: ATOMIC SPECTROSCOPY

Purpose:

**Part I:** Verify that Bohr’s theory on the structure of the H atom is correct, and determine the electron transition corresponding to each spectral line in the H emission spectrum.

**Part II:** Identify metal ions in two unknown aqueous solutions based on a flame test.

Performance Goals:

- Observe emission spectrum of hydrogen
- Prepare a calibration plot based on the helium emission spectrum
- Predict wavelengths of spectral lines in the H emission spectrum based on Bohr’s theory.
- Assign the spectral lines in the H emission spectrum
- Use a flame test to identify the cations in two unknown salts

Introduction:

Each element that makes up the natural world has its own unique signature in the form of the electromagnetic (EM) radiation that it absorbs or emits. This signature is referred to as the absorption or emission spectrum of the element. So unique is this signature that astronomers can determine the chemical composition of distant stars by analyzing the light they emit.

Early attempts to construct a theory that would explain and predict spectral properties began by considering the element hydrogen. Atoms of hydrogen are the simplest atoms that exist in nature consisting of only one proton and one electron and it seemed prudent to deal first with simple systems. In 1900 a German physicist by the name of Max Planck published a revolutionary paper showing that the energy associated with EM radiation could be determined from the frequency or wavelength of the radiation using the following equation.

\[ E = h \nu \]

where

- \( E \) = energy of EM radiation
- \( h \) = Planck’s constant = \( 6.626 \times 10^{-34} \) J\cdot s
- \( \nu \) = frequency of EM radiation in units of s\(^{-1}\) (cycles per second)
  or Hz (Hertz)

The frequency of the radiation is inversely proportional to the wavelength and can be given by the following equation.

\[ \nu = \frac{c}{\lambda} \]

where

- \( c \) = speed of light = \( 2.998 \times 10^8 \) m\cdot s\(^{-1}\)
- \( \lambda \) = wavelength of EM radiation in units of m

Substituting for the frequency, \( \nu \), in Equation 1 gives the following equation.

\[ E = \frac{hc}{\lambda} \]

Equation 2
Planck’s discovery set the stage for the first successful attempt to explain and predict the spectral properties of hydrogen. This was accomplished by Niels Bohr in 1913. Bohr pictured the hydrogen atom as a miniature solar system in which the tiny electron orbits about the much larger proton. According to Bohr, electrons were restricted to occupying only certain allowed orbits, each of which was denoted by a whole number ranging from 1 to \( \infty \) (infinity). These numbers were called principal quantum numbers. Bohr showed how the energy of an electron in a particular orbit could be calculated from the principal quantum number \( n \) of the orbit.

\[
E_n = \frac{-2.180 \times 10^{-18}}{n^2} \text{ J}
\]

Equation 3

where \( n \) = principal quantum number (ranging 1 to \( \infty \))

\( E_n \) = energy of electron for a particular quantum number, \( n \)

The electron orbit where \( n = 1 \) is the smallest orbit and has the least energy. When the single electron of hydrogen is in this orbit, the atom is said to be in its ground state, being in its most stable state. This electron can jump to orbits of higher energy where \( n = 2, 3, 4 \) or higher if it absorbs sufficient energy from some external source. Electrons are then said to be in an excited state. Electrons in excited states tend to return to the ground state and emit energy in the form of EM radiation when they do. The amount of energy emitted or gained by an electron (\( \Delta E \)) when it undergoes a jump from one allowed energy state to another can be calculated by first determining the energy of each state using Equation 3 and then finding the difference in energy (\( \Delta E \)) between the two energy states. Since emission is always due to the transition from a higher state to a lower state, \( \Delta E \) for such transitions is always \( E \) of the higher state minus \( E \) of the lower state, and it will always be a positive number.

\[
\Delta E = E_{\text{high}} - E_{\text{low}}
\]

Equation 4

This amount of energy (\( \Delta E \)) released is equal to the energy (\( E \)) of a photon of the light emitted (that is, \( \Delta E = E_{\text{photon}} \)). Max Planck had previously established an equation (Equation 2) showing the relationship between the energy (\( E \)) of a photon and the wavelength (\( \lambda \)), and frequency (\( \nu \)) of the light. Putting Equations 2 and 4 together Bohr was able to predict the wavelength and frequency of the light emitted for each transition of the electron of hydrogen as it jumps from one level to another.

\[
\Delta E_{\text{electron}} = E_{\text{photon}} = h \nu = \frac{hc}{\lambda}
\]

Note: Usually the subscripts are dropped and the equation becomes simply

\[
\Delta E = E = h \nu = \frac{hc}{\lambda}
\]

Equation 5

This equation, when rearranged to solve for \( \nu \) or \( \lambda \), becomes

\[
\nu = \frac{\Delta E}{h}
\]

or

\[
\lambda = \frac{hc}{\Delta E}
\]

Equation 5
As indicated by Equation 5, the wavelength ($\lambda$) of EM radiation is inversely proportional to its energy. Short waves have high energy and long waves have low energy. The human eye is only capable of seeing a narrow region (ranging in wavelength from 400 nm to 750 nm) of EM radiation called visible light. Visible light is made up of light of different colors but the human eye sees the light as white because it cannot separate the colors. If visible light is passed through a prism or diffraction grating it produces a continuous band of colors (red, orange, yellow, green, blue, indigo, and violet). But, when atoms are excited (in a gas discharge tube or in a flame), and the light emitted is passed through a prism or diffraction grating, it does not produce a continuous band of colors. Instead, a spectrum with discrete lines called a line spectrum or emission line spectrum is produced. Each element has its own unique set of EM radiation absorption and emission pattern, because the energy difference between one energy level to another ($\Delta E$) is different for each element, giving rise to lines with different colors being produced. For example, when sodium compounds are placed in a flame they produce a bright yellow flame while strontium compounds give brilliant reds. This is how we get the spectacular display of colors in fireworks. By examining the emission pattern from the stars we can determine what elements are there without traveling to the stars to obtain a sample.

In Part I of this experiment you will measure the wavelengths of the light emitted by hydrogen in the visible region of the EM spectrum and compare them with those calculated from Bohr’s theory. If the percent error is small we can conclude that Bohr’s theory is supported by experimental evidence.

Unfortunately, the observed wavelengths using our spectroscopes may not be accurate, but they can be corrected by first preparing a calibration curve using the known and observed wavelengths of the helium atom as a guide. The calibration curve is done by recording the observed wavelengths of the spectral lines of helium and plotting the known helium wavelengths (y-axis) against the observed helium wavelengths (x-axis). Each spectroscope will produce its own calibration curve. It is therefore important that you use the same spectroscope for both the hydrogen and helium emission spectra.

Once you have prepared the calibration curve, you will correct the observed values for hydrogen by reading them off your calibration curve. By comparing each corrected observed wavelengths of hydrogen with those predicted by Bohr’s theory, you will then identify each spectral line with the transition that produced it and verify whether Bohr’s theory is indeed valid.

In Part II, two unknown metal ions will be identified by comparing the color of light they emit with those of selected known metals ions using a flame test.

**Equipment/Materials:**
Spectroscope, electrodes receptacle, helium gas discharge tube, hydrogen gas discharge tube, known and unknown salt solutions, de-ionized water, gas burner, gas striker, a ceramic spotting plate, a piece of copper wire attached to a cork, wire cutter, and a 100- or 150-mL beaker
Procedure:

Part I: Spectral Lines of Helium & Hydrogen

A. Preparation of a Calibration Curve for the Spectroscope Using the Helium Spectrum

First check to make sure the power is off, then place a helium gas discharge tube in the receptacles of the light source. Turn on the light source. The helium tube should now be illuminated. Adjust the height of the spectroscope or the light source (with wooden boards or whatever is provided to be used) so that the spectroscope is positioned at the middle of the helium tube where the light is brightest. While looking through the eyepiece, aim the spectroscope at the light source by slowly moving it back and forth horizontally until a bright beam of light enters the slit at the front of the unit. The position should be fine-tuned so that this light beam is aimed straight at your eye. To the right of the slit, you will observe a virtual image of the helium spectrum displayed on a calibrated scale.

The scale you will see is shown below. Take a moment to first examine the scale, which is in nanometers (nm) (same as millimicrons, μm), and determine how many sig. fig. can be read from this scale. Record in your data sheet the wavelength reading, color for each bright band of light observed. It will make the analysis easier later on if you record the wavelengths in numerical decreasing order in “Table 10.1: Helium Emission Data.” That means start with the line on the far right and work towards the left. Turn off the light source when you are finished.

| 450 | 500 | 550 nm |

How should you record the positions of the two arrows in this figure? How many sig.fig. should they have? Check your answer with your instructor before proceeding.

You will see more lines than are listed in Table 10.1. Pick the brightest lines with colors that match. You must have a wavelength for each color line in the table.

B. Observation of Hydrogen Spectrum

With the power off, remove the helium tube from the light source and substitute a hydrogen tube.

CAUTION! THE HELIUM TUBE WILL BE HOT! USE FOLDED PAPER TOWELS TO HANDLE IT.

Turn on the light source and follow the same procedure as in Part A in observing and recording the hydrogen spectrum. (Remember that it is important that the same spectroscope be used for all wavelength measurements on both H₂ and He.) You should be able to see at least three lines. If you see a fourth line, record its color and
wavelength as well. Again, you will find it wise to do this systematically in numerical decreasing wavelengths in “Table 10.2: Hydrogen Emission Data.”

Analysis of Data in Part I:

A. Analysis of the Helium Spectral Lines
The accepted wavelength values and colors for helium are listed in Table 10.4. Next to each accepted value write in the observed wavelength for the value that most closely matches it. Make sure the color matches also. BEFORE YOU BEGIN PREPARING YOUR GRAPH, TABLE 10.4 MUST BE COMPLETED!

Preparation of the Calibration Graph Using Excel 2010

The objective is to prepare a graph using helium spectral lines to convert observed wavelengths to corrected wavelengths. The instructions below utilize Excel 2010, which is the version used on all CCBC college computers. If laptops are provided, you are expected to complete this graph in the lab. The instructions are brief as you have already been introduced to the use of Excel 2010 in an earlier experiment.

We begin with formatting the first two cells in Row 1 so that all content of each cell is visible within the cell by displaying it to multiple lines (Wrap text). The objective is to minimize the width of the data table so that there is more space for the graph in the final printout.

1. Click on the Home tab.
2. Highlight Row 1, Cells A and B. Then click on Wrap text.
3. In Row 1, label Column A Observed Wavelength (nm) and label Column B Corrected Wavelength (nm).
4. Enter in Column A, the Observed Wavelengths of Helium (from Table 10.4).
   Enter in Column B, the corresponding Accepted Wavelengths of Helium (from Table 10.4).
5. After all the data has been entered, highlight both columns, including Row 1, down to the last row that contains an entry.
6. Click on the Insert tab, on Scatter, and then on Scatter with only Markers.
Before proceeding, examine your graph. If there are any “outliers” first check to
make sure it is not due to an error in entering the data. If it is due to some other error,
it is most likely you matched your observed helium wavelengths with the wrong
accepted helium wavelengths. Get help from your instructor. You cannot proceed
until this is resolved.

7. Place your cursor anywhere on the spreadsheet (not on the graph) and click on Page
Layout tab, Orientation and select Landscape. The dotted lines that appear
indicate the size of the page you will be printing. Click on the frame of the graph and
then resize the graph so that it is as large as possible without letting it spill
beyond the dotted lines.
8. Add the line of best fit by clicking Layout, Trendline, and selecting Linear Trendline.
9. To obtain the equation for the trendline, click again on Layout, Trendline, and
select More Trendline Options at the very bottom. The Format Trendline window
will then appear.
10. Place a check mark at Display Equation on chart, and click on Close.
11. If necessary, move the equation to a position where it can be read easily (such as just
below the title). This is done by clicking on the equation once and then dragging it
to the desired position.
12. Next remove the legend (as it does not apply to this graph) by clicking on Legend
and selecting None.
13. Change the title by first clicking (only once) on the default title and then type in your
own title: Calibration Graph for Spectroscope: Corrected vs. Observed Wavelengths.
When you press ENTER, your title will appear on the graph. Alternatively, you can
highlight the default title and type in your own title.
14. Next click on Axis Titles, then select Primary Horizontal Axis Title, and Title
Below Axis. Type in a title for the x-axis: Observed Wavelengths (nm). Press
ENTER.
15. Click on Axis Titles, and then select Primary Vertical Axis Title, and Rotated
Title. Type in a title for the y-axis: Corrected Wavelengths (nm) Press ENTER.
As you have probably noticed, the data are bunched up in a small area of the graph.
We will now adjust the minimum and maximum for the x- and y-scales.
16. Click on Axes, select Primary Horizontal Axis and then on More Primary
Horizontal Axis Options. This is where you will adjust for the x-scale (Observed
Wavelengths). Select Fixed to allow you to change the default settings.
17. Remember the minimum must be smaller than your smallest x-value, and the
maximum must be larger than your largest x-value. Select Fixed to allow you to
change the default settings. If your lowest x-value is 440, and the maximum, around
660, try setting the minimum = 400 and maximum = 700. This may not work for
your data. You must choose settings that fit your data. Adjust the Major Unit and
Minor Unit as needed. Do not forget to go down to Minor tick mark type and
change “None” to “Outside.” When you are finished, click on Close.
18. Click on Axes, select Primary Vertical Axis and then on More Primary Vertical
Axis Options. Remember to select Fixed to change the default settings. Enter your
Minimum, maximum, Major unit, Minor unit. Go down to Minor tick mark type and change “None” to “Outside.” When you are finished, click on Close.

19. Click on Gridlines, select Primary Horizontal Gridlines and Minor Gridlines.
20. Click on Gridlines, select Primary Vertical Gridlines and Minor Gridlines.

21. In the area above the graph (such as in Cells F1 and F2), enter your course number, course section, semester, and names of students who prepared the graph. For example: CHEM 122 - Sec CM2 - Spring 2006 Jane Smith and John Doe
(It is important to include your names so that somebody else does not pick up your printout!) Drag and/or resize the graph if necessary in order to make space for the above information.

22. Highlight the entire page. Click on Page Layout tab, Print Area, and select Set Print Area.
23. Click on File, Print, and then Print Preview to double check that everything fits on the page, then click on Print. (Print 2 copies, one for each partner.)

24. Both partners must copy the equation of the trendline into his/her lab notebook.

B. Analysis of the Hydrogen Spectral Lines

1. Determining Correct Wavelengths for the Hydrogen Spectral Lines
   The Observed H Wavelengths in Table 10.2 are x-values. From the equation of the trendline, calculate y for each of the Observed H wavelengths and record the answers into Table 10.2 as Corrected H Wavelengths. These values should have the same significant figures as your observed wavelengths. You cannot have more precision in calculated values than in observed values. (Hint: These Corrected H Wavelengths should be in the general ballpark as the Observed H Wavelengths. In other words, if your Observed value is 650 nm, the Corrected value cannot be 1870 nm. If they differ significantly, check your calculations!) Show your calculation setups on a separate piece of paper to be submitted.

   The mechanism by which atoms produce light involves electrons undergoing transitions from higher quantum levels to lower ones. A transition from \( n = 7 \) to \( n = 2 \) should be entered as \( n = 7 \rightarrow 2 \), with the arrow showing the direction of the transition.

   Copy the “Corrected Observed H Wavelength” from Table 10.2 into Table 10.5. Next, match each of these corrected hydrogen wavelengths to the calculated hydrogen wavelength (from Appendix 4) that is closest in value to it. Record in the next column, the matching calculated hydrogen wavelengths. In the next column, write down the corresponding transition (e.g. \( n = 5 \rightarrow 2 \), or \( n = 4 \rightarrow 3 \) etc.).
3. Analysis of the Accuracy of the Observed H Lines.
   Calculate the error and % error of each of the “Corrected Observed H Wavelengths” using the theoretical wavelength provided in Appendix 4 as the accepted value. Refer to Experiment 2 on how to calculate error and % error. Enter the values in Table 10.5. What information does the % error give you?

Part II: Identification of Unknowns Using a Flame Test

A. Flame test for metal cations

   WEAR YOUR SAFETY GOGGLES AT ALL TIMES!

1. Your instructor will assign you two unknown samples in small vials, each containing one unknown metal cation in aqueous solution. Record the code number of the unknowns onto your Data Sheet.

2. Clean and dry a spot plate thoroughly then place it on a horizontal surface near where the gas burner will be used and add 4 to 6 drops of each of the following solutions to separate cavities on the plate: 1M NaCl(aq), 1M KCl(aq), 1M BaCl$_2$(aq), 1M Ca(NO$_3$)$_2$(aq), 1M LiCl(aq), 1M CuCl$_2$(aq), 1M NiCl$_2$(aq), 1M SrCl$_2$(aq). Make a diagram on a separate sheet of paper showing which solution is contained in each cavity.

3. Clean the copper wire thoroughly under the tap water, and then rinse it with deionized water. Fill the beaker with deionized water. This water will be used to rinse the copper wire between samples. Tap water cannot be used as it contains sodium ions which burn with a bright yellow-orange color and will mask any other colors that may be emitted by your test samples.

4. Light the gas burner and adjust the flame to give a hot intense blue flame about 12 cm high (about 5 inches) with an inner blue cone approximately 5 to 6 cm high (about 2 inches high). **If you have long hair, remember to tie your hair back!**

5. Begin by testing your copper wire. Hold it by the cork as the wire will get very hot. Place the tip (only) of the wire into the hottest part of the flame for about 3 or 4 seconds. Note the color of the flame as the wire glows red hot. The flame should show only a very pale yellow color. If it gives a green or crimson red color flame, it is contaminated. Hold it longer in the flame until the color is gone. If the color persists, ask your instructor to cut off the tip of the wire and test again.

6. When you are sure that the copper wire is clean, dip the tip in one of the known solutions, and again place the tip in the hottest part of the flame for about 3 or 4 seconds. Record the color of the flame, and whether it is fleeting or persistent.
Especially when two cations have similar colors, you will have to distinguish between them by noting whether the color disappears quickly (said to be “fleeting”) or lasts a relatively long time (said to be “persistent”).

7. Before you test your next sample, the wire must be cleaned so that it does not contaminate your sample. Heat the tip of the wire until the previous sample has been burned off completely (until the flame no longer shows the color of the previous sample). Allow the wire to cool down for about a minute, and then swish it around in the beaker of deionized water. Heat the tip of the wire again to ascertain that the wire is indeed clean (burning with a flame with no other color than the pale yellow). IF YOU ARE UNABLE TO GET RID OF THE PREVIOUS SAMPLE IN THIS MANNER CONSULT WITH YOUR INSTRUCTOR.

Test SrCl$_2$ last as it tends to stick to the wire and is hard to clean off. Before you go on to your unknowns, have your instructor snip off the end of the wire so that you can start with a fresh section of the wire.

Identify the unknown metal ions by comparing the flame color with those of the known metal ions. Remember to clean the wire between each flame tests as you had done for the known solutions. You will be graded on your accuracy of identification, so repeat this procedure until you are satisfied with your conclusion. (HINT: DO YOU HAVE ANY EVIDENCE FOR THE IDENTITY OF THE ANION?)

8. Write into Table 10.3, the formula of the metal ion that you conclude is present in each of your unknowns.

Prepare your Data Tables in your lab notebook.

PART I.

Table 10.1: Helium Emission Data

<table>
<thead>
<tr>
<th>Observed Color</th>
<th>Observed Wavelength (nm)</th>
</tr>
</thead>
<tbody>
<tr>
<td>red</td>
<td></td>
</tr>
<tr>
<td>yellow</td>
<td></td>
</tr>
<tr>
<td>green</td>
<td></td>
</tr>
<tr>
<td>green</td>
<td></td>
</tr>
<tr>
<td>blue-green</td>
<td></td>
</tr>
<tr>
<td>blue-violet</td>
<td></td>
</tr>
</tbody>
</table>

Table 10.2: Hydrogen Emission Data

<table>
<thead>
<tr>
<th>Observed Color of H Line</th>
<th>Observed H Wavelength (nm)</th>
<th>Corrected H Wavelength (from the Excel Trendline Equation)</th>
</tr>
</thead>
</table>
Part II: Identification of Unknowns Using a Flame Test

Remember to record the code numbers of your unknown!

Table 10.3: Observations & Conclusions from Flame Test

<table>
<thead>
<tr>
<th>SOLUTIONS</th>
<th>METAL ION PRESENT</th>
<th>FLAME COLOR</th>
</tr>
</thead>
<tbody>
<tr>
<td>none (dry Q-tip)</td>
<td>none</td>
<td></td>
</tr>
<tr>
<td>1M NaCl</td>
<td>Na⁺</td>
<td></td>
</tr>
<tr>
<td>1M KCl</td>
<td>K⁺</td>
<td></td>
</tr>
<tr>
<td>1M SrCl₂</td>
<td>Sr²⁺</td>
<td></td>
</tr>
<tr>
<td>1M BaCl₂</td>
<td>Ba²⁺</td>
<td></td>
</tr>
<tr>
<td>1M Ca(NO₃)₂</td>
<td>Ca²⁺</td>
<td></td>
</tr>
<tr>
<td>1M LiCl</td>
<td>Li⁺</td>
<td></td>
</tr>
<tr>
<td>1M CuCl₂</td>
<td>Cu²⁺</td>
<td></td>
</tr>
<tr>
<td>1M NiCl₂</td>
<td>Ni²⁺</td>
<td></td>
</tr>
</tbody>
</table>

Code # of unknown = ________
Conclusion: Observation:

Pre-Lab Exercise:

Your instructor will indicate whether you are to hand in answers to these questions or have you take a pre-lab quiz with similar questions.

1. Calculate the energy of an electron of the hydrogen atom when it is in the energy levels indicated:
   a) n = 3   b) n = 6   c) n = ∞

2. Using Bohr’s Equation calculate the energy change experienced by an electron when it undergoes transitions between the following energy levels.
   a) n = 6 → 3   b) n = 6 → ∞

3. Using Bohr’s Theory (not Rydberg’s equation) calculate the wavelength, in units of nanometers, of the electromagnetic radiation emitted for the electron transition n = 6 → 3. (1 m = 10⁹ nm).

4. In what region of the electromagnetic spectrum (UV? microwave? etc.) would the electromagnetic radiation occur for the transition n = 6 → n = 3? (Hint: See the section on Electromagnetic Radiation in your chemistry text.)

5. In the flame test, why must Sr²⁺ be tested last?

6. In the flame test if two ions have similar colors what will help you distinguish one from the other?
EXPERIMENT 10: SPECTROSCOPY

ELECTROMAGNETIC RADIATION SPECTRUM

Wavelength in nanometers

<table>
<thead>
<tr>
<th>10^{-2}</th>
<th>10^0</th>
<th>10^2</th>
<th>10^4</th>
<th>10^6</th>
<th>10^8 nm</th>
</tr>
</thead>
<tbody>
<tr>
<td>gamma rays</td>
<td>X rays</td>
<td>Ultraviolet rays</td>
<td>Infrared rays</td>
<td>Microwaves</td>
<td>Radio</td>
</tr>
</tbody>
</table>

Visible

<table>
<thead>
<tr>
<th>300 nm</th>
<th>400 nm</th>
<th>500 nm</th>
<th>600 nm</th>
<th>700 nm</th>
</tr>
</thead>
<tbody>
<tr>
<td>violet</td>
<td>blue</td>
<td>green</td>
<td>yellow</td>
<td>orange</td>
</tr>
</tbody>
</table>

Post-Lab Questions: (Please type your answers.)

1. What was the goal in Part I of the experiment? Write a conclusion to this part of the experiment, showing how you have reached the goal specified. Be as specific as you can.

2. In your own words explain fully, as precisely as you can, why hydrogen is emitting a blue-green light in this experiment.

3. Consider how your observations would be different if your spectroscope were accurate:
   a) How would you expect the observed wavelengths of helium to be different from the observed wavelengths you have recorded in Table 10.4? Be specific in your answer.
   b) What would you expect the value of the slope of your calibration graph to be?
   c) What would you expect the value of the y-intercept of your calibration graph to be?
   d) What would the trendline equation be?

4. Look up in the Internet what is known as the Lyman Series.
   a) Which transitions do the lines in the Lyman Series correspond to?
   b) Would you expect the emission lines in the Lyman Series to be in the visible region? Explain fully how you came to your conclusion. Include the reference source from the Internet in the proper format, including the access date.

5. Explain what a flame test has to do with spectroscopy.

6. (For this question only, answers are to be handwritten rather than typed.) Consider the electron of the H atom that undergoes the transition of n = 9 → 6.
   a) How much energy is released? Show your calculations carefully. As usual, all setups must include units.
   b) What would be the wavelength (in nm) of the light emitted? Show your calculations carefully. As usual, all setups must include units.
Tables of Results (for Part I)

Name: _______________________
Lab Sec: ______  Partner’s Name: _______________________

Table 10.4: Analysis of the Helium Spectral Lines

<table>
<thead>
<tr>
<th>Accepted Wavelengths of Helium (nm)</th>
<th>Color</th>
<th>Observed Wavelength of Helium (nm)</th>
</tr>
</thead>
<tbody>
<tr>
<td>668</td>
<td>Red</td>
<td></td>
</tr>
<tr>
<td>588</td>
<td>Yellow</td>
<td></td>
</tr>
<tr>
<td>502</td>
<td>Green</td>
<td></td>
</tr>
<tr>
<td>492</td>
<td>Green</td>
<td></td>
</tr>
<tr>
<td>471</td>
<td>blue-green</td>
<td></td>
</tr>
<tr>
<td>447</td>
<td>blue-violet</td>
<td></td>
</tr>
</tbody>
</table>

Table 10.5: Analysis of the Hydrogen Spectral Lines

<table>
<thead>
<tr>
<th>Corrected* Observed Hydrogen Wavelengths (nm)</th>
<th>Theoretical Hydrogen Wavelengths (nm)</th>
<th>Electron Transition</th>
<th>Error*</th>
<th>% Error**</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
<td></td>
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</tr>
</tbody>
</table>

* Show calculations for Corrected Observed Hydrogen Wavelengths on a separated sheet of paper.

**Show calculations of error and % error here: Watch sig. fig.!