

Lab 1. Atomic Spectrum of Hydrogen

Prelab Assignment

Before coming to lab:

- Use the handout "[Lab Notebook Policy](#)" as a guide to complete the following sections of your report for this lab exercise *before* attending lab: Title and Date of Lab, Introduction, Materials/Methods and Data Tables. Ensure that the table of contents of your lab notebook is current.
- Read the lab thoroughly and answer the pre-lab questions that appear at the end of this lab exercise. Background information can be found in [sections 7.1 and 7.2](#) in your textbook
- Complete the Prelab questions on the last page and hand in at the start of lab.

Purpose

In this experiment you will observe the line spectrum of hydrogen, find each line's wavelength, and determine the energy level transition corresponding to each line. You will also observe the flame spectra of several metals, and use your observations to identify three unknown samples.

Introduction

When an electric current is passed through a sample of gas in a sealed tube the energy excites the electrons of the atoms, causing them to jump to higher energy levels, n_h . Then, as the electrons fall back from the higher to lower energy levels, n_L , that energy is emitted as light. When this light is passed through a prism the result is not a continuous spectrum, but a **line spectrum**. In other words, only certain wavelengths of light are observed. This observation strongly suggests that the energy absorbed/emitted by an electron is **quantized**, or restricted to certain discrete values. Each line in the spectrum corresponds to a particular electronic transition between discrete energy levels.

In this lab you will measure the wavelengths of four lines in the atomic spectrum of hydrogen. The only lines you will be able to observe are those of the Balmer series, those lines that fall in the visible region of the spectrum (i.e. wavelengths between 400 and 700 nm). The lines of the **Balmer series** are the lines for which n_{lower} is equal to 2 (i.e. $n_L = 2$). Other transitions show up in other regions of the electromagnetic spectrum. For example, all transitions with $n_L = 1$ fall in the **UV region** of the spectrum, while all transitions with $n_L = 3$ fall in the **infrared**. With different detection equipment we could observe those transitions as well.

Once you have determined the wavelength, λ , of each line you can use the **Rydberg equation** (below) to determine the transition that produced each line. In other words, given that $n_L = 2$ (true for all lines in the visible region), you can solve for n_{higher} , n_h , the energy level from which the "excited" electron came down from.

$$\frac{1}{\lambda} = 1.096776 \times 10^7 \text{ m}^{-1} \left(\frac{1}{n_L^2} - \frac{1}{n_H^2} \right)$$

It is important to note that the Rydberg equation only works for hydrogen. While all elements produce a unique line spectrum, the Rydberg equation can be used to calculate initial and final energy levels only for hydrogen. Unfortunately, the Rydberg equation does not work for atoms with more than one electron. This is due to the presence of additional electron-nucleus attractions and electron-electron repulsions that are not accounted for by the Rydberg equation. Nonetheless, you may

choose to view the line spectra of other elements when you are finished with hydrogen. However, no calculations will be required for these other elements.

In addition to observing and analyzing the line spectrum of hydrogen, **flame spectra** of several different ionic compounds containing metallic ions will be observed. When a **metal** or **metal ion** is heated in a flame, the electrons are excited and promoted to higher energy levels just as they are in the gas tubes (why don't the electrons of the nonmetals get excited to a higher energy level?). When excited electrons return to a lower energy level the energy absorbed is released as visible, infrared and/or ultraviolet light. Because each element has a unique electronic structure, the color that it produces is also unique. Thus the flame spectra can be used to identify an element. Not all substances produce an observable color when heated because visible light is only a small part of the spectrum.

Procedure (Work in teams of 2 students)

Because we have only a limited number of spectrosopes available to use this lab will require some patience. You will work in groups of 2 and take turns making your readings. Fortunately, once each spectroscope is calibrated the readings do not take long. While you are waiting you will be able to carry out the procedure for the flame tests in part II, below.

Part I - Atomic Spectra

The instructor will demonstrate the use of the spectroscope, and will change the gas tubes as necessary.

1. Obtain a *Project STAR* spectroscope. Calibrate the spectroscope by pointing it at one of the fluorescent lights in the room and then adjust the plastic scale until the green line is aligned with the scale at 546 nm. If you are unable to get the **green line on the 546 nm mark**, record how many nanometers the calibration is off and use this value to correct for all future wavelength measurements that you make. Take note that the light from a fluorescent tube produces a bright-line spectrum (due to mercury vapor in the tube) superimposed over a continuous spectrum that is caused by the white light from the light source.
2. Using the **hydrogen tube**, record the wavelength in nanometers of the violet, blue-violet, blue-green and red lines in the hydrogen spectrum. Sometimes the violet line is quite faint and difficult to see—can you think of a reason why? Each member of the team should record the data individually, and then, when all are finished, compare your readings and recheck any that are in doubt.
3. If you choose, you may observe the spectrum of other gases we have available, *e.g.*, Argon, Neon, Chlorine, or Helium—ask your instructor to change the gas discharge tube for you. No data or calculations are required for these gases.

Part II - Flame Tests

1. Obtain a small container of each of the following salts:
LiCl, KCl, BaCl₂•2H₂O, SrCl₂•2H₂O, CaCl₂•2H₂O, NaCl, & Three Unknowns
2. Obtain a nichrome wire and a bottle of 6 M HCl. **CAUTION! HCl is CAUSTIC!** Place some of the acid in a small test tube.
3. Clean the wire by dipping it in the HCl in the test tube, and heating it in the hottest part of a Bunsen burner flame. Continue heating until no color, or only a faint yellow, is observed.
4. Dampen the wire with deionized water, dip it into one of the salts, and hold the sample in the hottest part of the flame, the tip of the inner blue cone of the flame. Record the color of the flame. Take care that none of the salt falls into the barrel of the Bunsen burner.
5. Repeat this procedure with each of the remaining salts and the **three unknowns**. **Remember to clean the wire between salts.**

Analysis of Results

Part I - Atomic Spectra

1. Use the Rydberg equation to calculate the transition associated with each line. (*I.e.*, what is n_i for each of the spectral lines observed in the hydrogen spectrum?)
2. Calculate the energy of each spectral line in units of Joules.

$$E = \frac{hc}{\lambda} \quad h = \text{Planck's constant} = 6.626 \times 10^{-34} \text{ J}\cdot\text{s}; \quad c = \text{Speed of light} = 3.00 \times 10^8 \text{ m/s}$$

3. Review your calculations and see if they make sense to you. Which color of light should be associated with the highest energy transition? The lowest? Why? Do your results agree with your prediction?
4. The true values for the wavelength of each line in the hydrogen spectrum can be found in your text. Compare your experimental wavelength for each spectral line to the theoretical values, calculate the % error for each and comment on the accuracy: Are the errors random or systematic? Are the errors within the precision of the measuring device? Explain fully and comment on the most likely sources of error.
5. Often the violet spectral line is too faint to be seen with our spectroscopes. Comment why the violet line is so faint, while the red spectral line is the brightest.

Part II - Flame Test

1. Each of the unknown compounds contains one of the following metal ions:
Li⁺, K⁺, Na⁺, Ba⁺², Sr⁺², Ca⁺²
 By comparison with the flame colors of the known compounds, identify the metal ion present in each of the three unknowns

Lab Report Checklist

As indicated previously, be sure to include all your lab data and calculations in your report, according to the guidelines in “[Lab Notebook Policy](#)” handout. Below is a checklist of the major things that should appear in each of the 5 sections of your lab report.

- | | |
|------------------------------|---|
| Introduction | <ul style="list-style-type: none"> • Goal/purpose of the lab is stated clearly? • Includes summary of background information • Includes Rydberg equation |
| Materials and Methods | <ul style="list-style-type: none"> • Procedures for parts 1 and 2 are detailed enough that a competent student could use it to replicate the experiment? |
| Results | <ul style="list-style-type: none"> • Ruled and easy to read data table(s) with the tables numbered consecutively and each with an informative caption? • All or most data summarized in one table for each part of the lab? <ul style="list-style-type: none"> — Data table for part 1 includes the measured and theoretical λ's, transition, energy and % Error for each spectral line? • Correct use of Sig figs and units? |
| Analysis of Results | <ul style="list-style-type: none"> • Work is clearly labeled with headings and is neat and easy to follow? • Calculation of electronic transition for each spectral line • Calculation of the energy of each spectral line in Joules • Comparison of results (λ's) with theoretical λ's and comments fully on the accuracy & sources of error. • Calculates % Error and discusses <u>possible</u> sources of error? • Correctly explains why the violet spectral line is very faint and the red line is very bright? • Correctly identifies the 3 unknown salts w/ flame test? |
| Conclusion | <ul style="list-style-type: none"> • Uses “<i>bullets</i>” to state concisely the major conclusions: <ul style="list-style-type: none"> — Part 1: line colors, λ's, % Error, transitions, line energies — Part 2: flame colors of salts, ID of unknowns • Summarizes sources of error or uncertainty? |

Overall quality: easy to follow, neat and appropriate depth of discussion?

Therapeutic Doodle Space

Lab 1. Atomic Spectrum of Hydrogen
Prelab QuestionsName _____
Team # _____ Date _____ Section _____

Instructions: Complete the following four questions in the spaces provided below and hand in at the start of your lab period. Show your work with *units and correct significant figures* for all questions that involve a calculation. **Circle all numerical answers.**

1. How many possibilities exist for emission spectral lines when a hydrogen atom's electron in the $n = 4$ quantum level of a Bohr atom drops to the ground state (i.e. $n = 1$)? Enter your response in the space provided and explain your reasoning. *Use a diagram to illustrate each of the possible transitions responsible for the spectral lines. Your reasoning must be correct to receive credit.*

Number of spectral lines = _____

Reasoning:

2. Use the Rydberg equation to calculate the wavelength in nm of the light emitted when the electron in a hydrogen atom undergoes a transition from $n = 6$ to $n = 2$. Consult [figure 7.8](#) in your textbook to determine the color of light emitted. *Show your work with units and sig. figs. Circle your answer.*

3. Calculate the energy in units of Joules for each photon of light in the spectral line in question #2, above. **Show your work with units and sig. figs.** Circle your answer.

$$E = \frac{hc}{\lambda} \quad h = \text{Planck's constant} = 6.626 \times 10^{-34} \text{ J}\cdot\text{s}; \quad c = \text{Speed of light} = 3.00 \times 10^8 \text{ m/s}$$

4. The electron in a ground-state hydrogen atom (i.e. $n = 1$) absorbs a photon of wavelength 93.8 nm. Use the Rydberg equation to determine the energy level the electron moves to. **Hints:** (i.) The Rydberg equation gives the relationship between the wavelengths of light emitted when an electron within a hydrogen atom falls from a higher energy level, n_i , to a lower energy level, n_f . (ii.) The energy absorbed by an atom to excite an electron to a higher energy level (in this case a photon of $\lambda = 93.8$ nm) is equal to the energy emitted when the electron returns to the energy level from which it came (in this case, $n = 1$). Circle your answer.