Lab 2: Bohr Atom and Emission Spectroscopy

Objectives: To investigate the atomic structure of hydrogen atom and to learn the theory of absorption and emission of light by atoms.

Equipment: Emission spectrometer; emission tubes of hydrogen, helium, krypton, and water gases.

Safety: Use caution when handling the high voltage source used with the emission tubes. Wear safety goggles and gloves at all times.

Waste: none

INTRODUCTION

The electronic structure of elements is responsible for much of the chemical behavior of matter. Most of what we know about electronic structure is the result of **spectroscopy**, or the study of how matter interacts with electromagnetic radiation. Spectroscopy has provided experimental evidence for many of the theories concerning the energy and location of electrons in atoms. To fully appreciate how electromagnetic radiation interacts with matter we need to discuss the properties of electromagnetic radiation, or light.

Light can be described by both wave theory and particle theory. The wave-like properties of light include wavelength and frequency, which are related by Equation 2.1:

Equation 2.1. $c = \lambda \cdot \nu$

where c is the speed of light in vacuum (2.998×10^8 m/s), λ is the wavelength (in m), and ν is the frequency (in s^{-1}). But light can also be treated as discrete packets of energy, called photons. The energy of a photon is defined by Equation 2.2:

Equation 2.2. $E_{photon} = h \cdot \nu$

where E is the energy of the photon (in Joules) and h is Planck's constant (6.6266×10^{-34} J·s). Combining Equations 2.1 and 2.2 yields:

Equation 2.3. $E_{photon} = hc/\lambda$

Equation 2.3 shows the constant relationship between the energy of a photon and its wavelength.

A **spectrum** is a plot of the intensity of light as a function of wavelength. There are two types of spectra. In a **continuous spectrum,** the intensity of light varies smoothly as a function of wavelength. An example is sunlight. When passed through a prism, sunlight is dispersed into a continuous band of colors of relatively equal intensity. The different colors that we see are due to the different wavelengths of electromagnetic radiation present in the sunlight spectrum.

By contrast, a **line spectrum** contains only discrete wavelengths, or "lines" with dark regions in between. Early investigators noted that different elements would emit line spectra when excited either thermally or electrically. Each element exhibits a unique line spectrum or **emission**

spectrum that can be used to identify that element. For example, the emission spectrum of hydrogen atom contains four discrete lines at specific wavelengths.

John Blamer (1825-1898) showed that frequencies of the lines in hydrogen spectrum can be formulated into a mathematical equation. Balmer's equation was later expanded by Johannes Rydberg (1854-1919) to what is now called Rydberg's equation, Equation 2.4.

Equation 2.4. $$\frac{1}{\lambda} = R_H \left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

where λ is the wavelength of the spectral line (nm), R_H is the Rydberg's constant (1.096776×10^{-2} nm^{-1}), n_1 and n_2 are positive integers, $n_2 > n_1$. This equation can be used to calculate the wavelength of a light associated with transition from n_2 to n_1.

The first model to explain the line spectra of the elements was proposed by Niels Bohr in the early 1900s. This model can be summarized by a few fundamental assumptions:

1. Electrons are located in orbits that are found at specific distances from the nucleus.
2. The distance from the nucleus is related to the energy level of the orbit; the orbit closest to the nucleus has the lowest energy, and the energy increases with distance from the nucleus.
3. The energy levels for the orbits are **quantized**—they can only have specified energies that vary in a step-like fashion.

The energy associated with an electron in a given orbit can be described by Equation 2.5:

Equation 2.5. $$E_n = -B/n^2$$

where E_n is the energy of the n^th orbit and B is the Bohr constant (1312 kJ/mol).

Bohr postulated that the line spectra of the elements resulted from the movement of electrons between the quantized energy levels. Electrons can move to higher energy levels as a result of thermal or electric excitation. When the electron moves back down to a lower energy level, it emits a photon of light having a wavelength that corresponds to the difference in energy between the initial and final states (ΔE). Each spectral line that we observe is the result of the movement, or **transition**, of an electron between energy levels. Since the energy levels of the orbits are fixed, the energy associated with an electronic transition between two specified energy levels is also fixed and can be calculated by Equation 2.6:

Equation 2.6. $$\Delta E = E_{final} - E_{initial} = -B\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right)$$

where n_i and n_f are positive integers ($n_i > n_f$) representing the initial level (n_i =higher energy) and final level (n_f =lower energy) of the electron, respectively. Equation 2.6 implies that the *ΔE for emission will always be negative*, which simply means that energy is *released* rather than absorbed.

The energy diagram for electron transitions can be depicted as shown in Figure 2.1. Transitions to a higher energy level represents **absorption** of light and transitions to a lower energy level represents **emission** of light.

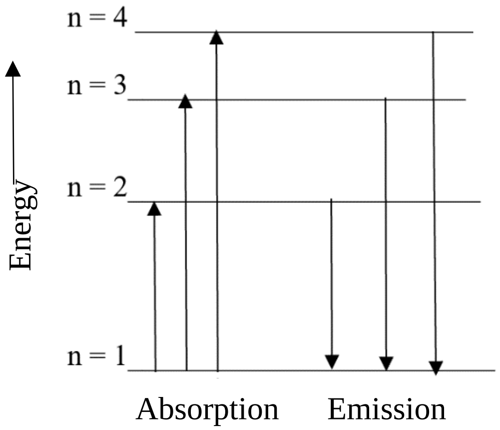


Figure 2.1. Example of an energy level diagram

Unfortunately, the electronic structure of atoms is more complex than the simple model proposed by Bohr. The Bohr model works well for explaining the line spectra for the hydrogen atom, which contains only a single electron, but the model represented by Equation 2.6 fails when applied to multi-electron atoms. In this lab you will use spectroscopy to evaluate the Bohr model for the hydrogen atom and to examine the line spectra of various elements.
In the first two parts of this experiment, you will record emission spectrum of hydrogen atom and calculate the energy, wavelength and energy level, n_i, of each transition in emission spectrum of hydrogen atom.

In the last part you will observe emission spectra of different gases and compare it to hydrogen spectrum.

Name: ______________________ *Bohr Atom and Emission Spectroscopy*
Section:____________________ *Date:_______________*

Pre-Lab Questions

1. Define each of the following terms and give an example to illustrate.

a. Line spectra –

b. Continuous spectra –

c. Quantized –

d. Transition –

2. Using the Bohr model, calculate the energy associated with the transition from n = 4 to n = 1 (in kJ/mol).

3. What is the wavelength of light (in nm) associated with the transition described in question?

4. Explain the difference between emission and absorption of energy.

PROCEDURE

Part A. Calculation of the Emission Lines for Hydrogen atom, Balmer Series

In the Bohr model, the Balmer series includes the lines emitted by transitions of the electron from a higher energy level ($n_i > 2$) to the 2nd energy level ($n_f = 2$) in hydrogen atom. Four of these emission lines have a wavelength in the visible light range and can be seen with naked eyes. In this part you will calculate the energy and wavelength of these four emission lines.

1. Calculate the energy differences (ΔE) for the emission transitions represented in the table using Equation 2.6. The units of ΔE will be in kJ/mol. Show your calculations to your TA.

2. To calculate the wavelength associated with each transition, rearrange Equations 2.1 and 2.2 to yield:

Equation 2.7. $\lambda = hc / E$

where h = Planck's constant (6.626×10^{-34} J·s), c is the speed of light (2.998×10^{8} m/s), and E is in the energy in J/photon (rather than kJ/mole). Using your ΔE value (in kJ/mol) and Equation 2.8 will yield a wavelength in nm:

Equation 2.8. $$\lambda\ (\text{nm}) = \frac{1.196 \times 10^{5}\ (\text{kJ.nm/mol})}{\Delta E\ (\text{kJ/mol})}$$

Record your values in the bottom half of the appropriate cells in Data Table 2.1. Show your calculations to your TA.

Part B. Emission Spectrum of Hydrogen

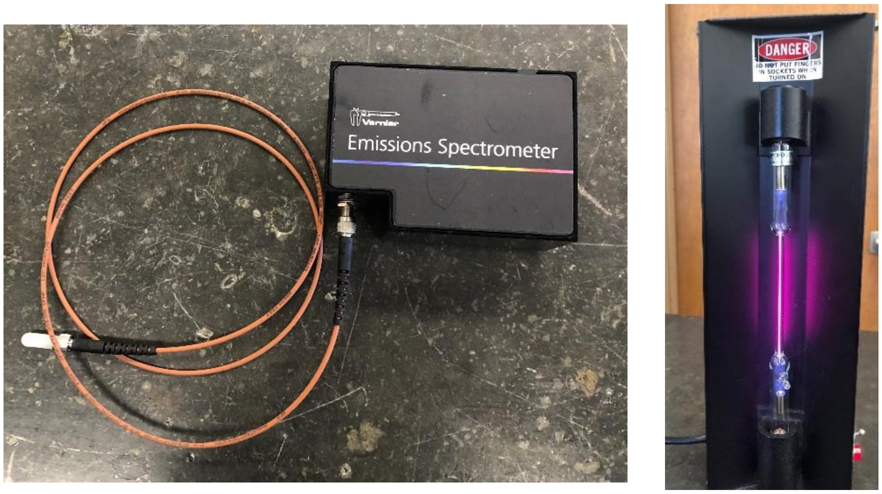
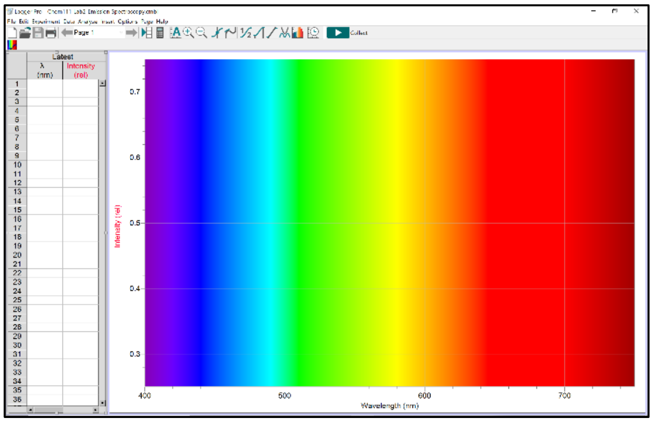


Figure 2.2. Vernier emission spectrometer, optical fiber, and hydrogen tube.

In this part, you will use Logger Pro software and Vernier emission to observe Balmer series line spectrum of hydrogen atom.
There are six emission lines in Balmer series for hydrogen but only four of them lie in the visible light range and can be detected by the spectrometer that we use.

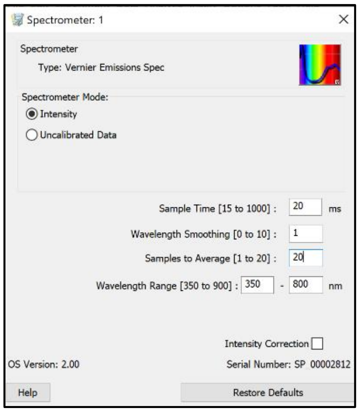
Preparing software:

3. Log in to the desktop using your NIU credentials and open Logger Pro and CHEM212-Lab2-Emission Spectrometry. You should see a data table and continues visible light spectrum (rainbow).



4. From **Experiment** menu, select **Set Up Sensors > Spectrometer 1** and enter the following values for data collection:
    a) Spectrometer mode: Intensity
    b) Sample Time: 20 ms
    c) Wavelength Smoothing: 1
    d) Samples to Average: 20
    e) Wavelength Range: 380 – 800 nm
    f) Close the window

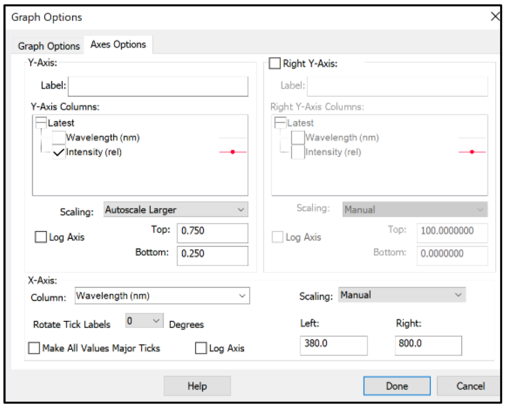
This will tell the software to collect 20 sets of data for 20 ms, average, smooth, and display the collected data in 350 – 800 nm wavelength range.



Spectrometer
Type: Vernier Emissions Spec
Spectrometer Mode:
● Intensity
○ Uncalibrated Data
Sample Time [15 to 1000] : 20 ms
Wavelength Smoothing [0 to 10] : 1
Samples to Average [1 to 20] : 20
Wavelength Range [350 to 900] : 350 - 800 nm
Intensity Correction ☐
OS Version: 2.00
Serial Number: SP 00002812

5. Right click anywhere on the graph and select **Graph Options**
    a) In the **Title** box, enter the name of emission tube that you are using.
    b) Click on the **Axes Options** tab, choose **Autoscale Larger** from **Scaling** in **Y-axis** area.
    c) From Scaling in **X-axis** area, choose **Manual**, and enter 380 in left box and 800 in the right box.
    d) Close the window.

This will set the scale for X and Y axes on the graph.



Y-Axis:
Label:
Y-Axis Columns:
Latest
☐ Wavelength (nm)
☑ Intensity (rel)
Scaling: Autoscale Larger
☐ Log Axis
Top: 0.750
Bottom: 0.250
☐ Right Y-Axis:
Label:
Right Y-Axis Columns:
Latest
☐ Wavelength (nm)
☐ Intensity (rel)
Scaling: Manual
☐ Log Axis
Top: 100.0000000
Bottom: 0.0000000
X-Axis:
Column: Wavelength (nm)
Scaling: Manual
Rotate Tick Labels 0 Degrees
Left: 380.0
Right: 800.0
☐ Make All Values Major Ticks
☐ Log Axis

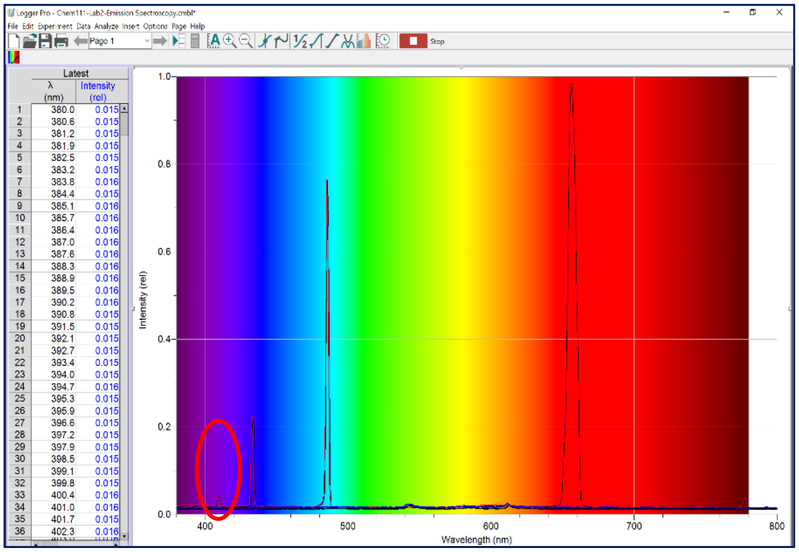
Data Collection:

6. Place the hydrogen tube in the power supply and turn on. You need to wait 2-3 minutes for the tube to warm up. Be careful, tube will be hot!

7. Remove the protective cap from the spectrometer optical fiber and save it in a small beaker. You must put the cap back on after finishing the data collection to avoid damaging the optical fiber.

8. Gently hold the probe close to the middle of hydrogen tube. Hold steady and do not let the probe touch the tube. Tube is hot!

9. Click **Collect** to start the data collection.
   - a) You may need to move the probe back and forth a little to adjust the intensity of peaks.
   - b) Each time you move the probe, it will take some time for display to refresh.
   - c) Click **Stop** when the four peaks of Balmer series are displayed with a good intensity.
   - d) The emission line in purple region has a low intensity, as indicated in Figure 2.3,and it may take few tries to record it.



	Latest	
	λ (nm)	Intensity (rel)
1	380.0	0.015
2	380.6	0.015
3	381.2	0.015
4	381.9	0.015
5	382.5	0.015
6	383.2	0.015
7	383.8	0.016
8	384.4	0.015
9	385.1	0.016
10	385.7	0.016
11	386.4	0.016
12	387.0	0.016
13	387.6	0.016
14	388.3	0.016
15	388.9	0.016
16	389.5	0.016
17	390.2	0.016
18	390.8	0.015
19	391.5	0.015
20	392.1	0.015
21	392.7	0.015
22	393.4	0.015
23	394.0	0.015
24	394.7	0.016
25	395.3	0.015
26	395.9	0.015
27	396.6	0.015
28	397.2	0.015
29	397.9	0.015
30	398.5	0.015
31	399.1	0.015
32	399.8	0.015
33	400.4	0.016
34	401.0	0.016
35	401.7	0.015
36	402.3	0.016

Figure 2.3. Balmer emission lines of hydrogen atom.

10. You may repeat the data collection, if you are not satisfied with your emission spectrum or cannot see all four peaks. From **Experiment** menu, choose **Clear the Latest Run** and repeat step 9. Below are some tips for changing the intensity so that all lines fit into the scale.
    - a) Moving the optical fiber further away from the tube decreases the intensity.
    - b) Moving the optical fiber closer to the tube increases the intensity.
    - c) Waiting longer increases the intensity of smaller peaks.

11. From the **Experiment** menu, select **Store Latest Run**, after you have a good quality emission spectrum.

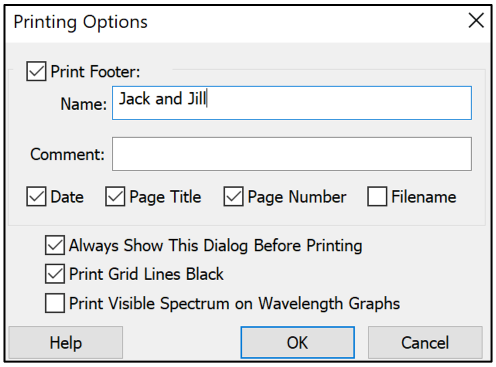
12. Turn off the power supply and let the hydrogen tube to cool down.

13. To determine the wavelength of each emission line, move the cursor to the top of one peak, note the set of numbers on the lower left corner of the graph. The first number is the wavelength (λ) in

nanometers (nm). The second number, separated by a comma, is the intensity of the peak. Record the wavelength of each peak in Data Table 2.2.

Saving your graph

14. From **File** menu, select **Print Graph**.
    a) Check **Print Footer** and type the names of lab partners.
    b) Check the boxes as shown in **Printing Options**. Click **OK**.
    c) Take a screenshot of the graph and save on desktop.
    d) Email a copy of the graph to each lab partner. You need to upload this graph along with your data sheet.



☑ Print Footer:
Name: Jack and Jill
Comment:
☑ Date ☑ Page Title ☑ Page Number ☐ Filename
☑ Always Show This Dialog Before Printing
☑ Print Grid Lines Black
☐ Print Visible Spectrum on Wavelength Graphs

Do not exit the software.

Part C. Emission Spectrum of Helium and Krypton

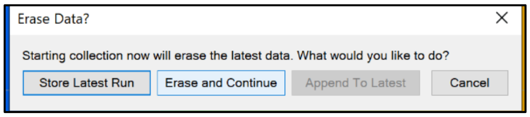
These atoms have more electrons than hydrogen, so their emission spectra is more complex than hydrogen with only 2 electrons.

15. Remove the hydrogen tube from the power supply. It must be cool by now.

16. Place helium tube in the power supply, turn on and wait 2-3 minutes for the tube to warm up.

17. Repeat steps 4 and 5 to set up experiment's parameter.

18. Click **Collect**, and **Erase and Continue**. Wait until you have peaks with good intensity. **Click Stop**.



Starting collection now will erase the latest data. What would you like to do?

19. Turn off the power supply.

20. Record the wavelength and color of each emission line in Data Table 2.3.

21. Save a copy of the helium emission spectrum. You need to upload a copy of this graph along with your data sheet.

22. Place the krypton tube in the power supply, collect the spectrum, and record your observations in Data Table 2.3.

23. You may try different tubes to observe their emission spectra.

24. Turn off the power supply and return the tubes to the cart.

25. Replace the fiber optic's cap. Exit Logger Pro. Do not save.

CALCULATIONS

Part A. Calculation of the Emission Lines for Hydrogen atom, Balmer Series

1. Calculate ΔE using Equation 2.6.

2. Calculate λ using Equation 2.8.

Part B. Data Collection of Emission Lines for Hydrogen, Balmer Series

3. Use your observed wavelength (λ) and Rydberg's equation to calculate the initial energy level (n_i) for each emission line. In Ballmer series $n_f = 2$ and $R_H = 1.096776 \text{ x } 10^{-2} \text{ nm}^{-1}$ so Rydberg's can be simplified as:

Equation 2.9. $$\frac{1}{\lambda} = -1.096776 \text{ x } 10^{-2} \text{ nm}^{-1} \left(\frac{1}{2^2} - \frac{1}{n_i^2}\right)$$

4. Convert your observed wavelength (λ) to units of *m* and use Equation 2.3 to calculate energy of one photon, E_{photon} (J) for each emission line.

5. Calculate energy of mole of photons for each emission line in kJ/mol.

Equation 2.10.

$$E_{one\ mole\ photon}\ (kJ/mol) = E_{photon}(\text{J/photon}) \times \frac{1\ kJ}{1000\ J} \times \frac{6.0221 \times 10^{23} photon}{1\ mol}$$

Name: ______________________ *Bohr Atom and Emission Spectroscopy*
Section: ____________________ *Date:*_____________

Data Sheet

Data Table 2.1. Balmer Series Data for the Hydrogen Atom

Energy Level	n_i=3	n_i =4	n_i =5	n=6
n_f =2	Δ*E* (kJ/mol) =	Δ*E*(kJ/mol) =	Δ*E*(kJ/mol) =	Δ*E*(kJ/mol) =
	λ(nm)=	λ(nm)=	λ(nm)=	λ(nm)=

Data Table 2.2. Hydrogen Spectrum

	Wavelength (λ) in nm	n_f	n_i	E_{photon} (*J*)	E one mole photon (*kJ/mol*)
1st peak (Start from right) Red region		2			
2nd peak Light blue-green region		2			
3rd peak Blue-purple region		2			
4th peak Dark purple region		2			

Name: ______________________ *Bohr Atom and Emission Spectroscopy*

Section: ____________________ *Date:______________*

Data Table 2.3. Helium and Krypton Spectra

Helium Color of the tube:	**Peaks**	
	Wavelength (nm)	Color of the Background
Krypton Color of the tube:		

Name: ______________________ *Bohr Atom and Emission Spectroscopy*
Section: ____________________ *Date:*______________

Post-Lab Questions

1. In Part A, you calculated the wavelength of four lines in the Balmer series of the hydrogen atom. In Part B, you experimentally determined the wavelengths of four lines. Is the calculated and experimental data consistent? If not, what might be causing the discrepancy?

2. Based on your observations, what is the correlation between electron configuration of an atom and its emission spectrum?

3. The emission spectrum of a nebula shows emission lines at wavelengths 640 and 649 nm. Based on your observations, what gas may be present in this nebula?

4. If you were to use a tube filled with mixture of gasses, would the observed spectra be the sum of the spectra for the individual gasses or the average of the individual spectra? Explain.

5. a) Calculate ΔE(kJ/mol) and λ(nm) for transition from n=5 to n=3.

   b) Compare your answer to ΔE and λvalues for transition from n=5 to n=2 in hydrogen Balmer series. Comment on relationship between energy levels, ΔE, and λ.