Spectroscopy and the Hydrogen Atom

Purpose
In this experiment, you will predict the wavelengths of light emitted by hydrogen atoms by using calculations from the Bohr model of hydrogen. Then you will use a simple spectroscope to observe the light emitted by hydrogen gas discharge tubes and other sources.

Background
Light is electromagnetic radiation, composed of oscillating electric and magnetic fields that are perpendicular to each other. The speed of electromagnetic radiation is given by the constant c (by definition exactly 2.99792458×10^8 m/s in a vacuum, for convenience 2.998×10^8 m/s). Light is often described by waves, which are periodic disturbances through space.

Light has two important properties: 1) amplitude, which gives the maximum extent of the oscillation and is related to the intensity (or brightness) of the light; and 2) the wavelength of light which is the distance from two corresponding points along a wave.

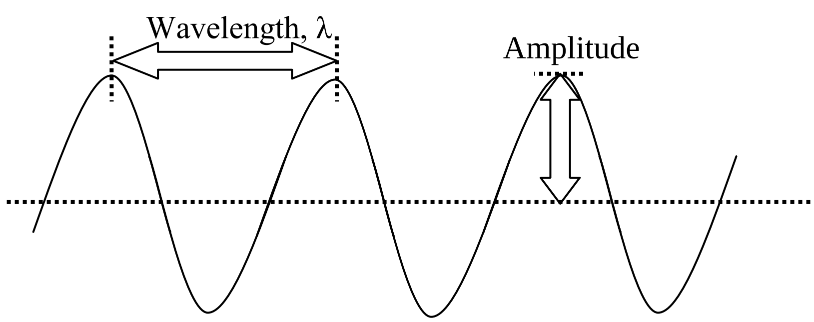


Figure 1: Electromagnetic Wave Characteristics

The wavelength of light is related to the frequency of the wave by Equation 1. The frequency of the wave is the number of complete wavelengths that pass a given point each second, and is given in units of inverse seconds: s^{-1} (or 1/s, also called a Hertz (Hz)).

$$\upsilon \cdot \lambda = c \qquad \textbf{(1)}$$

where υ is the frequency, λ is the wavelength, and c is the speed of light.

The wavelength, frequency, and energy of light has the relationship:

$$E = h\upsilon \quad \text{or} \quad E = \frac{hc}{\lambda} \qquad \textbf{(2)}$$

where h is Planck's constant (6.626×10^{-34} J·s)

The color of light observed is due to the wavelength (and corresponding frequency or energy) of the light.

Transfer the answers for Q1-Q7 which appear in the following pages to the prelab assignment to turn in before lab.

Q1) What do the following symbols represent? What are the SI units of each? E, υ, λ, h, c

Q2) Use the relationships in Equation 1 and 2 (above) to fill in the blank with "increases" or "decreases"

a) **As λ increases, υ _________**
b) **b) As λ increases, E _________**
c) **c) As υ decreases, E _________**

Q3) Red light has a wavelength of 700 nm. Violet light has a wavelength of 400 nm. Which photons of light are higher in energy, red or violet?

In the early 20^{th} century, it was accepted that light also had particle-like properties. Light as a stream of particles (called photons) rather than a wave explains some observations, but not all. Therefore, light is regarded as both wave-like and particle-like. This is called the wave-particle duality of light. It was later discovered that electrons also exhibit a dual nature of wave and particle.

Spectroscopy and the Bohr Model

Atomic structure is quite strange. At the end of the 19^{th} century, evidence about the structure of atoms had scientists wondering how atoms could even exist at all. Physics predicted that the electron should spiral into the nucleus, emitting light of all colors (frequencies) until it collapsed. The contradiction was that atoms seemed to exist, and they emitted only certain colors of light. Classical physics could not explain this phenomenon. A new physics, called quantum mechanics, was created to formulate theories about the nature of very small particles, which behaved in mysterious ways that the physics of the day could not explain.

In 1913, Danish physicist Niels Bohr proposed a model of the atom that was consistent with the observations. To solve the problem of electrons spiraling into the nucleus, he suggested that electrons were confined to certain "orbits" (energy levels). They could move from one orbit to another, but could not stay in between orbits or go to the nucleus.

Here's how the model works: The atom in its lowest energy configuration (the electron occupies the lowest energy orbit) is called the ground state configuration. (Figure 2)

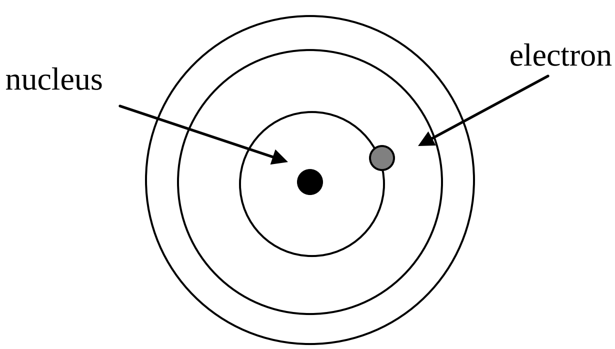


Figure 2: Ground state (lower energy) configuration

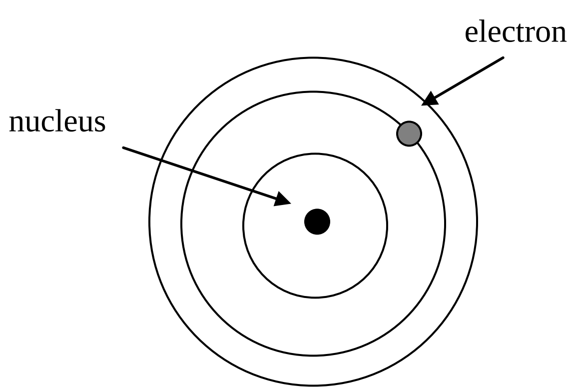


Figure 3: Excited state (higher energy) configuration

When atoms absorb energy (from heat, light, or electricity), the electrons in the orbits may have enough energy to "jump" to higher levels. These atoms are in an "excited" state configuration (see Figure 3). As these excited atoms relax and fall to lower levels, they release the energy (as light). The energy of the photons of light that are released depends on the initial and final states of the atom. Electrons that "fall" the greatest distance will release more energy than those that "fall" a shorter distance.

Picture two situations: One in which an electron falls from the 4th energy level to the 3rd, and another in which an electron falls from the 4th energy level to the 2nd.

Q4) Which transition (4→3 or 4→2) will emit higher energy photons?

Q5) Which transition (4→3 or 4→2) will emit photons of a longer wavelength?

Q6) Will an electron absorb energy or release energy to undergo a 1→4 transition?

Since electrons are restricted to certain (quantized) energy levels, the amounts of energy they release are discrete (non-continuous). **<u>Quantized</u>** energy levels explain why only certain colors of light are observed rather than a continuous spectrum of color.

Q7) Find the definition of "Quantum" in your textbook/on the internet. Summarize the definition in your own words.

Bohr's model is attractive for other reasons as well. Because of the electrons "orbiting" a nucleus many referred to this model as a "planetary model" of the atom. How elegant it seems to have the motion of very tiny particles described by the same model as our solar system! This feeling did not last long. Bohr's model has some shortcomings (for example, it cannot explain observed spectra for systems that involve more than one electron). Nonetheless, it is a historically important step toward the modern atomic theory.

Procedure

Part 1 of this lab involves many calculations. Unless otherwise stated by your instructor, you will do these calculations in lab.

First, the energy levels of the hydrogen atom will be calculated and plotted on an energy-level diagram. Next, the energy changes for electron transitions between those levels, as well as the wavelengths of photons corresponding with those transitions will be calculated. Finally, these calculated values will be used to assign transitions to experimentally observed wavelengths.

Part 2 of this lab involves observing the spectra of incandescent light and fluorescent light, as well hydrogen and other elements.

Data

Unless otherwise directed by your instructor, it is recommended that in Part 1 you reproduce the data tables and the energy level diagram shown below in your lab notebook. For Part 2, set up space to record the observed spectra in your lab notebook.

Part 1: Predicting the visible spectrum of hydrogen

A. Calculations of the energy levels of the hydrogen atom

Each energy level in the atom has a specific value of energy. It is given by the following equation:

$$E_n = -\frac{Z^2(2.179\times10^{-18}\,\mathrm{J})}{n^2} \qquad (\text{where } n = 1, 2, 3, \ldots\ ) \qquad (3)$$

where n is the energy level, E_n is the energy of level n, and Z is the atomic number of the element.

For hydrogen, Z=1 and equation 3 becomes $$E_n = \frac{-2.179\times10^{-18}\,\mathrm{J\ per\ atom}}{n^2} \qquad (4)$$

1) Equation 4 above gives the energy per hydrogen atom. The energies will be easier to work with if we consider a mole of atoms. Fill in the equation below, converting the constant given (-2.179x10^{-18} J/atom) to units of kJ per mole. **Check your answer with other groups and your instructor before you proceed!** (Hint: If the energy is -2.179 × 10^{-18} J for one atom, will the energy become bigger or smaller for a mole of atoms? More negative, or less negative?)

$$E_n = \frac{\underline{\qquad\qquad\qquad\qquad}\ \mathrm{kJ\ per\ mole}}{n^2} \qquad (5)$$

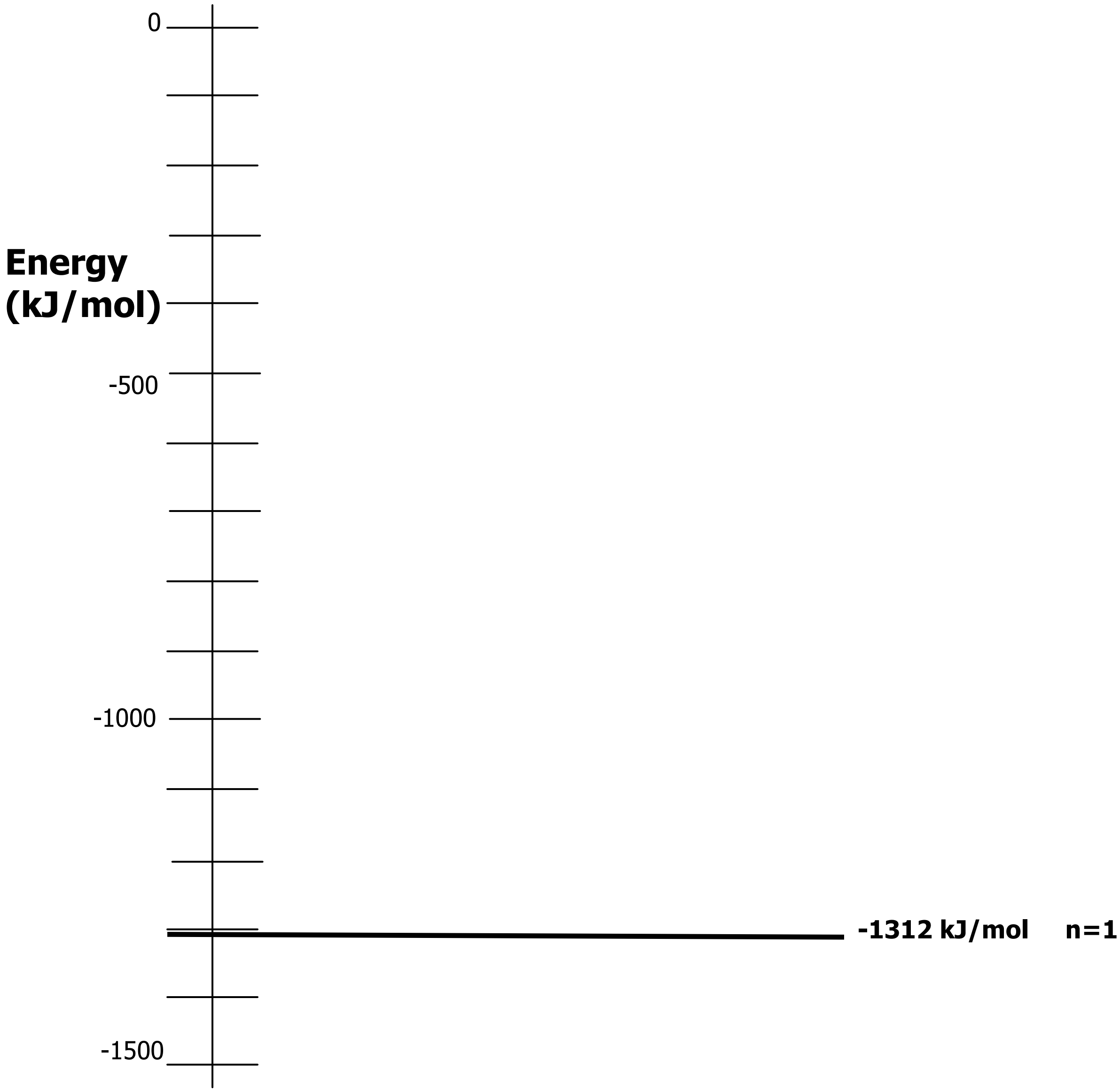
2) Theoretically, an atom has an infinite number of energy levels. However, we will focus on the first ten. Using the equation you determined in Question 1 above (giving kJ per mole of atoms), complete the following table. These energies correspond to the energy of an electron in a particular energy level.

Table 1

Quantum Number, n	Energy, E_n (kJ/mol)	Quantum Number, n	Energy, E_n (kJ/mol)
1		6	
2		7	
3		8	
4		9	
5		10	

3) Plot the energy values in Table 1 on the graph below by drawing a horizontal line across the y-axis corresponding to the energy value for each of the 10 levels. Write the energy level on the right end of the horizontal line. The first one (n = 1) is done for you as an example. You will soon see that it is difficult to get all the lines in the graph—just do the best you can!

Energy-level Diagram for Hydrogen



4) What do you notice about the spacing of the energy levels as n increases?

5) What is the value of E_n when $n = \infty$? Put this into your graph.

B. Calculation of the wavelengths of the lines in the hydrogen spectrum

The observed light emitted by hydrogen is due to the transitions of electrons in hydrogen atoms from higher energy levels to lower energy levels. Therefore, the energy of the photons is calculated by

$$\mathbf{E_{photon} = |\Delta E| = |E_{final} - E_{initial}|}$$

<u>To complete the upper half of each box</u>, use the values of E calculated in Table 1 to find the change in energy using the equation above. For example, the ΔE for an electron going from the second energy level to the first is -328 – (-1312) = 984. Do this for all the upper halves of the boxes. The lower half of each box is explained below.

<u>To complete the lower half of each box</u>, calculate the wavelengths associated with the energy changes that were calculated in the upper half of the box.

Use the relationship: $$\Delta E = \frac{N_A hc}{\lambda(\text{nm})} \qquad (5)$$

which rearranges to

$$\lambda = \frac{N_A hc}{\Delta E} \qquad (6)$$

where N_A = Avogadro's constant, h = Planck's constant, and c = speed of light.

Legend for Table 2

Note: Unit conversions must be performed to obtain ΔE and λ in the desired units.

top box	ΔE (kJ/mol)
bottom box	λ (nm)

Table 2

n (lower) \ n (higher)	6	5	4	3	2	1
1					984	
					122	
2						
3						
4						
5						

C. Assignment of electronic transitions to wavelengths

Using Table 2 from Part B, assign as many of the following electron transitions as you can. The first is shown as an example: 97.25 nm corresponds with an electron transition from n =4 to n = 1.

The wavelengths may not exactly match your calculated ones. There will be four boxes left blank—these will be explained with Table 4.

Table 3

Wavelength (nm)	Transition $n_{higher} \rightarrow n_{lower}$	Wavelength (nm)	Transition $n_{higher} \rightarrow n_{lower}$	Wavelength (nm)	Transition $n_{higher} \rightarrow n_{lower}$
97.25	$4 \rightarrow 1$	410.29		1005.2	
102.57		434.17		1094.1	
121.57		486.27		1282.2	
389.02		656.47		1875.6	
397.12		954.86		4052.3	

There were four unassigned wavelengths above. Transfer the values of these unassigned wavelengths to Table 4.

They are unassigned because Table 2 only considers up to six energy levels. You calculated the energies of the first ten (see Table 1). Using Table 1 and Table 2, take a guess at its transition. Then do some calculations to see if you are correct. The first one is done for you (without the calculation shown).

Table 4

Unassigned wavelengths (nm) (from Table 3)	ΔE (kJ/mol)	Probable transition ($n_{higher} \rightarrow n_{lower}$)
389.02	307.5	$8 \rightarrow 2$

<u>Part 2: Observing the visible spectrum of hydrogen and other elements.</u>

Obtain a spectroscope and practice looking through it. First use a spectroscope to view an incandescent light bulb. You should see a rainbow of colors. Adjust the spectroscope until you see a wide band of color. If you are having difficulty, ask your instructor for help.

Now use the spectroscope to view the fluorescent lights in the ceiling. You should see a bright green line and a bright violet line (along with a few others).

Draw the **<u>lines</u>** that you see using colored pencils:

Incandescent light

Fluorescent light

The light that is emitted from gases at low pressure results in **<u>line</u>** spectra (as opposed to continuous spectra). Several gas discharge tubes may be set up for viewing.

For each of the following gases, sketch the **<u>LINES</u>** that you see with the colored pencils corresponding to the colors that you see. Start with hydrogen.

Name of element: Hydrogen

Estimate the values of the wavelengths of the lines you observe! Do this especially for hydrogen, using the measuring device in the spectroscope.

Remember, visible light is between 400-700 nm.

Name of element: Helium

Name of element: Neon

Name of element: Mercury

Name of element: _______________

Report Sheets Name__________________ Date________
Spectroscopy Lab partner______________Section____

Part 1: Predicting the Visible Spectrum of Hydrogen - The Balmer Series

In 1885, a Swiss schoolteacher named Johann Balmer noticed four lines in the spectrum of hydrogen. These four lines are part of what is now called the Balmer Series. At the time, Balmer was limited to detecting the visible and near-ultraviolet wavelengths, but it is now possible to detect wavelengths in other regions of the electromagnetic spectrum.

1a) List the electron transitions in Table 3 that are part of the Balmer Series.

1b) List the electron transitions in Table 4 that are also part of the Balmer Series.

1c) Besides being in the visible (or near visible) range, what common characteristic do the lines in the Balmer Series have *with respect to the assignments you gave them for their electronic transitions?*

Use ΔE and the idea that energy levels are discrete (quantized) in order to answer the following questions:

2a) Considering an infinite number of energy levels in an atom, **calculate** the **shortest** possible wavelength for a line in the Balmer Series. HINT: Does the shortest possible wavelength correspond to the smallest or largest ΔE value possible? What transition is required?

2b) Why can't shorter wavelengths than this one be possible in the Balmer Series?
Think about the energy levels in an atom.

3a) Considering an infinite number of energy levels in an atom, **calculate** the **longest** possible wavelength for a line in the Balmer Series. HINT: Does the longest possible wavelength correspond to the smallest or largest ΔE? What transition is required?

3b) Why can't longer wavelengths than this one be possible in the Balmer Series?
(The answer is not "because it is outside the visible and near-ultraviolet ranges". Think about the energy levels in an atom. Look at Table 2 and the energy diagram.)

Part 2: Observing the Visible Spectrum of Hydrogen and Other Elements

1) In Part 1 of this lab you identified the electron transitions for hydrogen that are part of the Balmer Series. Based on this, how many lines of the Balmer Series in the hydrogen spectrum **should be** observed using the spectroscope? ________

2) How many of the lines in the Balmer Series for hydrogen were you actually able to observe? (Look at your drawing for hydrogen in Part 2). List the specific values of the wavelengths (in nm) that you were able to observe. ________

3) As you may have noticed, each element has its own unique line spectrum. Line spectra can be used for identification (for example, astronomers can use spectroscopy to determine the elemental composition of stars).

Fluorescent lights are also gas discharge tubes, as you saw in this experiment. They contain several elements, one of which was in this lab. If you were able to view other gases (such as helium, neon, mercury, nitrogen, etc…) you may be able to identify one of the elements in fluorescent lights. **Which gas from this experiment is in fluorescent bulbs based on your observations?** (It may not be identical due to the presence of other elements. Which is the most probable, based on what you observed?) **Explain**.

4) Summarize what you learned from this lab by writing a short paragraph.

Guiding questions to address:
a) What is the Bohr Model? How does it work? What evidence supported the model?
b) What is the meaning of quantization of energy?
c) Even though the Bohr Model was later replaced by another model of the atom, why is it important to know the Bohr model?

Answer these questions on a separate sheet, or typed (ask your instructor). USE YOUR OWN WORDS (do not cut and paste from the internet).

A high quality response will be factually correct, have a logical flow, and include the following terms: orbits, energy levels, quantized energy, absorption and emission of energy as light, electron transitions, line spectra. **Underline or highlight these terms in your paragraph.**

Pre-Lab Assignment
Spectroscopy and the hydrogen atom

For more information the concepts in this lab, refer to your textbook.

DIRECTIONS: Read and answer the questions in the background section of this lab. These questions are reproduced below so you can turn this sheet into your instructor at the beginning of lab.

1. What do the following symbols represent? What are the units of each?

E

υ

λ

h

c

2. Use the relationships in Equation 1 and 2 (in the previous section) to fill in the blank with "increases" or "decreases"

a) As λ increases, υ __________

b) As λ increases, E __________

c) As υ decreases, E __________

3. **Circle** your choices for the following questions.

a) Red light has a wavelength of 700 nm. Violet light has a wavelength of 400 nm.

Which color of light (**red** or **violet**) is higher in energy?

b) Which transition (**4→3** or **4→2**) will emit higher energy photons?

c) Which transition (**4→3** or **4→2**) will emit photons of a longer wavelength?

d) Will an electron (**absorb** or **release**) energy to undergo a 1→4 transition?

4. Find the definition of "quantum" in your textbook/on the internet. Summarize the definition in your own words. Cite your source.