

THE BOHR MODEL AND SPECTROSCOPY OF THE HYDROGEN ATOM

OBJECTIVES

To perform calculations associated with the Bohr model of the hydrogen atom and compare the results to the observed spectrum of the hydrogen atom.

MATERIALS

Solutions of metal salts

EQUIPMENT

Spectrometer; emission tubes of various elements

SAFETY

Use caution when handling the high voltage source used with the emission tubes. Wear safety goggles when working with the Bunsen burner and metal salt solutions.

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 wavelike properties of light include wavelength and frequency, which are related by Equation (1)

$$c = \lambda \cdot \nu \quad (1)$$

where c is the speed of light in vacuum ($2.998 \cdot 10^8$ 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)

$$E = h \cdot \nu \quad (2)$$

where E is the energy of the photon (in Joules), and h is Planck's constant (6.6266×10^{-34} J·s). Combining Eqs. (1) and (2) yields:

$$E = \frac{hc}{\lambda} \quad (3)$$

Eq. (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 spectra**, 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 that can be used to identify that element.

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 (4).

$$E_n = -\frac{B}{n^2} \quad (4)$$

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. 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 (5)

$$\Delta E = E_{final} - E_{initial} = \left(\frac{1}{n_f^2} - \frac{1}{n_i^2} \right) \quad (5)$$

where n_i and n_f represent the initial (higher energy) level, and final (lower energy) level of the electron, respectively. Eq. (5) implies that the ΔE will always be negative, which simply means that energy is released rather than absorbed.

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 Eq. (5) 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.

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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 $n = 3$ to $n = 1$ (in kJ/mol).

3. What is the wavelength of light (in nm) associated with the transition described in Question 2?

PROCEDURE

Part A: Calculation of the Emission Lines for Hydrogen

1. Calculate the energies of the first ten orbits (i.e., for $n = 1$ through $n = 10$) for the hydrogen atom using Eq. (4). Record your results in the first column of Table 1 in Part A on the Data Sheet. Round each energy to the nearest 0.1 kJ/mol.
2. In columns 2 through 11 calculate the energy differences (ΔE) for the emission transitions represented in the table using Eq. (5). The units of ΔE will be in kJ/mol. As an example, the ΔE value for the $n = 4$ to the $n = 2$ level has been calculated and entered in the table for you. Check your calculations against the value provided. Enter your calculated ΔE values in the top half of the appropriate cells in the table.
3. Once the ΔE values for all transitions have been calculated, calculate the wavelength associated with each transition by rearranging Eqs. (1) and (2) to yield

$$\lambda = \frac{hc}{E} \quad (6)$$

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 (7) will yield a wavelength in nm.

$$\lambda = \frac{1/196 \times 10^5 \text{ kJ} \cdot \text{nm}}{\Delta E \text{ mol}} \quad (7)$$

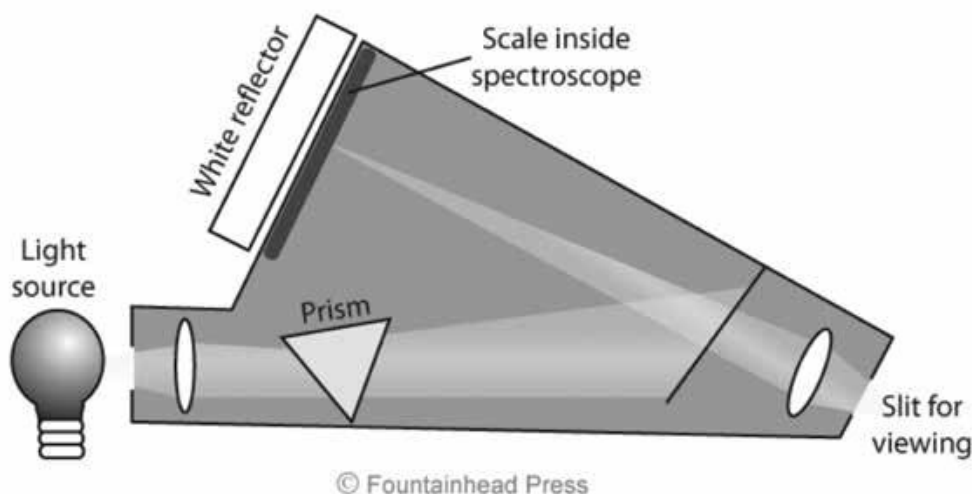
Record your values in the bottom half of the appropriate cells in Table 1. Again, the wavelength for the $n = 4$ to $n = 2$ transition has been calculated for you. Check your calculations against this value.

Part B: Calibration of the Spectroscope

4. A diagram of a typical spectroscope is provided in Figure 1. Light from the discharge lamps enters the spectroscope through a narrow slit on the left side in the diagram and is diffracted into separate bands of color. These bands are then focused on a white reflector which indicates the approximate wavelength of each band, or spectral line (in nm). It is important that the spectrometer results are calibrated by checking the measured wavelengths against the known wavelengths for these lines. Once you have calibrated your spectroscope you should use the same instrument for all your lab measurements.

View the mercury emission lamp through your spectroscope. Record the color and position (in nm) of each line that you observe on Part B of the Data Sheet. Using the actual mercury transition wavelengths provided, create a calibration curve of nm (observed) vs. nm (actual) for the mercury emission line data. Using this graph, you can convert the nm reading for the observed emission lines for other elements into the actual wavelength in nm.

Figure 1. Diagram of a typical spectroscope.



Part C: Hydrogen Spectra

5. Follow the directions provided in Part C of the Data Sheet.

Part D: Emission Lines

6. View the emission lines for the emission tubes provided. For each tube, record the name of the element, the color of each emission line, and the observed wavelength (in nm) in Part D of the Data Sheet.

Part E: Flame Spectra

7. Set up a Bunsen burner so that the flame is near the entrance slit of the spectroscope. Working in pairs, observe the flame emission spectra for the metal salt solutions provided. One student should monitor the emission spectra while his/her partner introduces samples of the salt solution in the Bunsen burner flame. Dip the metal loop into the salt solution and hold the loop in the flame. You may have to do this several times to observe and record all the emission line data for a given salt solution. Students should then switch places and repeat the procedure for another salt solution. After all the salt solutions have been tested, partners should share all their data so each student has a complete set of emission data for the metal salt solutions. Wear safety goggles when working with the salt solutions and Bunsen burner.
8. Obtain an unknown salt solution from your instructor and record the emission lines observed for your unknown following the procedure in Part E step 7. Compare the results for your unknown with the emission spectra for the known solutions. Identify the unknown.

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DATA SHEET 1

Part A. Calculation of the Emission lines for hydrogen

Calculate the energies of the first ten orbits for the hydrogen atom and enter them in the first column of Table 1. Calculate the ΔE values and wavelengths (nm) for each of the transitions in Table 1 and record these values in the appropriate cell locations in the table.

Energies (kJ/mol)		Transition Data								
		n=2	n=3	n=4	n=5	n=6	n=7	n=8	n=9	n=10
n=1; E=?	$\Delta E =$									
	$\lambda =$									
n=2	$\Delta E =$									
	$\lambda =$									
n=3	$\Delta E =$									
	$\lambda =$									
n=4	$\Delta E =$									
	$\lambda =$									
n=5	$\Delta E =$									
	$\lambda =$									
n=6	$\Delta E =$									
	$\lambda =$									
n=7	$\Delta E =$									
	$\lambda =$									
n=8	$\Delta E =$									
	$\lambda =$									
n=9	$\Delta E =$									
	$\lambda =$									
n=10	$\Delta E =$									
	$\lambda =$									

Part B: Calibration of the Spectroscope

		Actual	Observed
Mercury Lines:	faint violet	366 nm	_____nm
	bright violet	405 nm	_____nm
	faint blue	436 nm	_____nm
	double green	546 nm	_____nm
	fuzzy yellow:	580 nm	_____nm
	thick red	615 nm	_____nm

Part C: Hydrogen Spectra

View the hydrogen emission spectrum using the hydrogen lamp and the spectroscope. In the space below, draw the lines you observe and their relative spacing. Indicate the colors and approximate wavelengths of each line.

Label the lines in the picture above as A, B, C and D. Using the transition data from Table 1, identify the transitions for each line.

A =

B =

C =

D =

Part D: Emission Lines

Element	Line Color	Location (nm)

Part E: Flame Spectra

Salt	Formula	Lines: Color/Wavelengths
Unknown ID:		

POST-LAB QUESTIONS

1. Why is it necessary to prepare a calibration plot in Part B?
2. If you were to use a solution of mixed salts in Part E, would the observed spectra be the sum of the spectra for the individual metal salts or the average of the individual spectra? Explain.
3. The wavelengths for the emission lines observed in the mercury lamp spectra do not match the wavelengths calculated using the Bohr model in Part A for the hydrogen atom. Explain why.