Atomic Models and Spectroscopy

**In this activity you will use both a computer simulation and laboratory instruments to investigate the line emission spectra for different atoms.**

**Introduction**

When light from the Sun or white light is passed through a prism it produces a “rainbow” of different colors. This has been known for centuries, with Isaac Newton making significant advances in the area of “Opticks” in the 1700’s by investigating how light is reflected, refracted, dispersed, etc.

However, something very different occurs when light from a single element is emitted and passed through a prism. In this case only a small number of discrete lines are observed. Even more remarkably, the pattern of these lines (the **line spectrum**) is a defining characteristic of each element. What happens when light is emitted from an element like hydrogen? Why is the pattern different for different elements? As you will see in this investigation, the answering of these question has led to profound insights into the electronic structure of the atom and, in turn, our modern understanding of chemistry.

**Logistics**

This laboratory activity is designed for one 3-hour session for General Chemistry students working in groups of 2-3. Parts A.1 and A.2 may be completed prior to attending lab. The PhET sim, if used as a standalone activity in Part A, can be completed in about 60-90 minutes. Prior knowledge regarding the structure of the atom and atomic models proposed by Dalton, Thomson, and Rutherford is assumed and extended here.

**Part A. Investigate Models of the Hydrogen Atom**

Descriptions of different models of the hydrogen atom are scattered throughout your textbook. Atomic models are conceptual. A weather map is also an example of a conceptual model. The weather map communicates some (but not all) aspects of the weather, e.g. the temperature, areas of high and low pressure, precipitation, etc. Models allow one to make predictions (will it rain in New York tomorrow?) and makes sense of complex systems. **Atomic models are conceptual** and serve to communicate different characteristics and make predictions, even when it is impossible to directly see the parts of the atom itself.

**A.1) Use the sim to collect the line emission spectrum for hydrogen.**

Start by playing with the sim “Models of the Hydrogen Atom**”**

Figure out how to do the following things with the sim:

* Turn on the “gun” to direct light into a box of hydrogen.
* Speed up, slow down, or pause the sim.
* Use a spectrometer to detect photons emitted from the box.
* Direct only green light (photons of “green”) in to the box of hydrogen.
* Direct only ultra-violet (UV) photons in to the box of hydrogen.
* Collect the line emission spectrum using the spectrometer when white light enters the box. It may be helpful to speed up the sim. After a few minutes **sketch the spectrum in the box below[[1]](#footnote-1).** Include in your sketch the *colors (regions of the electromagnetic spectrum)* and estimated *wavelengths*.

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* Also figure out how to take a picture of the spectrum. Save the spectrum on the screen. This is “experimental data” and can be compared with predicted line spectra for different atomic models.

**For hydrogen, how many emission lines are present in the visible light region, and which color do you think will be the brightest, i.e. the most intense?**

**A.2) Investigate the pre-quantum models of the hydrogen atom.**

* Figure out how to switch to different atomic models in the sim. This will allow you to test different models and see the line spectrum they predict.

For the “pre-quantum” models (Dalton’s Billiard Ball model, Thomson’s Plum Pudding model, and Rutherford’s Classical Solar System model) perform the following tasks:

* In your notebook, describe each model with words and/or a sketch.
* Test the model: Does it produce a line emission spectrum matching the experimental one?

Dalton’s atomic model was put forward in 1807 and was influential for much of the 19th century. It is not uncommon for conceptual models to be replaced, but also retain value if they are still useful for describing some aspects of nature.

**Are there parts of Dalton’s atomic hypothesis still used by chemists today?**

Thomson’s model was proposed in the last years of the 19th century. Scientific ideas are often described as “concrete and durable”, with strong support from different sources, but at the same time “tentative”, with ideas being subject to change in light of new evidence or new interpretations of existing evidence.

**What led to Thomson’s model replacing Dalton’s model? Was it new evidence, or a new interpretation of existing evidence?**

Some textbooks suggest that Rutherford’s model was a great success and widely accepted upon its introduction. This is false, as it took several years before it became accepted as a plausible model.

**Why do you think the Classical Solar System model was not readily accepted?**

**A.3) Investigate the Bohr and de Broglie models of the hydrogen atom.**

First, figure out how to show the electron energy level diagram for these models.

There are now three different representations shown in the sim: the “picture” of the atom, the electronic energy level diagram, and the predicted emission spectrum. Your task is to make connections between all three of these domains.

In your notebook, use words and/or sketches to compare and contrast the Bohr and de Broglie models in terms of:

1. What the atom “looks like”,
2. The electron energy levels, and
3. The emission spectra.

**How is the location of the electron related to the energy level diagram?**

**What causes the electron to change its position?**

**What causes a photon to be emitted?**

Go to the *Help* menu, select *Transition*, and then **fill in the missing entries:**

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| --- | --- | --- |
| Electronic transition between n levels | Wavelength (nm) | UV, visible, or IR?(Look at the spectrometer) |
| 1 ⇔ 2 | 122 | UV |
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| 1 ⇔ 4 | 97 | UV |
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| 2 ⇔ 3 | 655 | Visible |
|  |  |  |
|  |  |  |
| 2 ⇔ 6 | 410 | Visible |
| 3 ⇔ 4 | 1876 | IR |
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| 3 ⇔ 6 | 1094 |  |
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| 4 ⇔ 6 |  | IR |
| 5 ⇔ 6 |  | IR |

In General Chemistry, common calculations involve converting between energy, frequency, and wavelength:

E=*h*ν Planck’s relationship for the energy of a photon

c=λν 1 nm = 1x10-9 m *h* = 6.63x10-34 Js.

c= 3.00x108 m/s in a vacuum.

You may also be familiar with the concept of “matter waves,” where the wavelength of any particle is inversely related to the particle’s mass and speed: λ = *h*/mv

**Suppose an electron is at the n=2 energy level. It is then struck by light with a frequency of 4.5770x1014 s-1. What happens to the electron?**

**In the sim’s representation of de Broglie’s model of the hydrogen, the wave-character of the electron and the proton appear quite different. Why do you think this is?**

Suppose the electron is in orbit closest to the nucleus. **Complete the table below** by FIRST predicting what will happen if monochromatic radiation is used. AFTER you have recorded your predictions, test them with the sim. It may help to perform tests after completing a row.

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| --- | --- | --- | --- |
| Wavelength | Movement of electron | Energy level diagram | Emission spectrum |
| 103 nm | Prediction: | Prediction: | Prediction: |
| 97 nm | Prediction: | Prediction: | Prediction: |
| 99 nm | Prediction: | Prediction: | Prediction: |

Were your predictions confirmed by the sim?

If so, you have a good understanding of these models.

**A.4) Investigate the Schrödinger model of the hydrogen atom.**

*Erwin Schrödinger began working on wave mechanics in 1925. His interest was sparked by a footnote in a paper by Albert Einstein. Like de Broglie, he began to think about explaining the movement of an electron in an atom as a wave. In 1926 he published his work, providing a theoretical basis for the atomic model that Bohr had proposed. The equation at the heart of his publication became known as Schrödinger's wave equation.*

In your notebook, use words and/or sketches to compare and contrast the Schrödinger and the Bohr models in terms of:

1. What the atom “looks like”,
2. The electron energy levels, and
3. The emission spectra.

Notice, as the electron moves in to different orbitals the quantum numbers change**. Use white light and, for three different combinations of quantum number in the sim, sketch the resulting atomic orbital in the table below.**

|  |  |  |
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| Quantum Numbers | Quantum Numbers | Quantum Numbers |
| Sketch of atomic orbital | Sketch of atomic orbital | Sketch of atomic orbital |

Let the sim run with white light and identify the electronictransitions (i.e. the change from one orbital to another) that take place. **Add your results to the table below**. An example for the absorption of a photon is given (how do you know it’s an absorption?)… it is *not*necessary to list the corresponding emission, e.g. from n,l,m = 4, 1, 0 to the ground state 1, 0, 0 as well.

**Electronic Transitions in the Hydrogen Atom**

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| Quantum State(Initial) | Quantum State(Final) | Does the transition involve the ground state? |
| n | L | m | n | L | m | Yes |
| 1 | 0 | 0 | 4 | 1 | 0 |
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There are many energy levels and corresponding quantum numbers for the hydrogen atom, but the electron does not move between every possible combination. In other words, some transitions “allowed” and others “forbidden”.

Start by looking at information in your table for transitions involving the ground state.

**Are there “rules” restricting the transition of the electron involving “n”, or can it change by any amount?**

 **Are there “rules” involving “*l*”?**

Now look at transitions between excited states.

**Do your rules explain why an electron in an excited state might absorb another photon and moves to a different excited state instead of returning directly to the ground state?**

**Part B.**  **Use a Spectroscope to Observe the Line Spectra for**

 **Hydrogen and Helium**

**In this activity you will…**

* + Record scale readings for hydrogen emission lines with a spectroscope.
	+ Prepare a calibration curve for the spectroscope.
	+ Record scale readings for helium emission lines with a spectroscope.
	+ Use the calibration curve to determine the wavelength of the He emission lines.

The line spectrum for hydrogen you will observe is what results when white light enters the ?-box in the computer simulation (although here the excitations take place with electricity). As shown in the simulation, photons with many different wavelengths (UV, visible, IR) are emitted. Photons in the visible light region are observable with our eyes, but we do not see a line spectrum because our eyes observe all of the photons at the same time. To produce a line spectrum the photons must be separated. This is accomplished with a **spectroscope**.

As you might infer from the image below, spectroscopes have been around for a long time. They make use of the phenomena of refraction; by simply passing light through a prism, the photons of different energy may be separated.

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As shown by visible light passing through a prism, the extent of the refraction (the change in direction) is affected by the wavelength.

**Are high-energy photons refracted more or less than low energy photons?**

**With a spectroscope you observe the “screen” following the dispersion of the light. Would the line spectrum you see change in any way if the distance between the prism and the screen changed?**

Based on the sim, how many lines do you expect to see for a hydrogen discharge lamp? For each line, predict the color, the wavelength, and whether it will be “bright” or “dim”. Base your predictions on what you observed using the sim and record the information in your notebook.

Color Wavelength Intensity Scale reading

 (nm) (bright of dim)

AFTER RECORDING YOUR PREDICTIONS, use the spectroscope and write the scale readings in your notebook for the lines you observe (see cautions below). These are the emission lines appearing on the “screen” of the spectroscope.

When using the spectroscope you will see a bright yellow line. This is from the element sodium, which is in the glass. It has an emission line at 589 nm. Add this value to your table as well.

**Cautions and Procedure**

* The 5000 V step-up transformer and contacts for tubes on boards are an electrical hazard.
* Always wear your lab goggles since they will block out almost all-harmful UV radiation.
* Make sure the overhead lights in the balance room are off during observations.
* The TA will assist you with setting up the spectroscope.

**Data Analysis for Hydrogen: Calibration of Spectroscope**

The spectroscope separated the photons of light emitted from the hydrogen discharge lamp based on their energy. Since you already know the wavelength values for the different emission lines you can **calibrate** the instrument by finding a relationship between wavelength and the scale readings. Plotting and then fitting the data with a program like Excel is a useful way to determine this relationship. Complete this task and record the calibration information in your notebook.

**Helium Emission Line Spectrum**

A helium discharge lamp is also available in lab. For this element you do not know in advance the number of emission lines or their wavelength. Fortunately, after your analysis of the hydrogen emission lines you do have a calibrated spectroscope. Use the spectroscope to identify the emission lines for helium and the accompanying scale readings. Then, use your calibration data to convert these scale readings to wavelengths. Record your data, calculations, etc., in your notebook.

Now **make a sketch of the line emission spectrum for helium in the box below**. Notice the x-axis has units of wavelength.

**He line spectrum: Data from a spectroscope**

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 Wavelength (nm)

**Part C.**  **Using a Spectrometer[[2]](#footnote-2) to Observe the Line Spectra**

Although simple, inexpensive, and historically interesting, spectroscopes are not frequently used by chemists. In contrast, **spectrometers** are extremely important. In fact, it is difficult to imagine modern scientific laboratories without numerous spectrometers. You will use several kinds of spectrometers this semester. How are they different from spectroscopes? To answer this question, use a spectrometer to record the emission from a helium discharge lamps and **sketch it in the box below**.

**He line spectrum: Data from a spectrometer.**

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 Wavelength (nm)

**Why might a chemist favor a spectrometer over a spectroscope?**

**Report Questions**

**The report questions are very important in this laboratory report. Convince the reader you understand the concepts investigated with the computer simulation and/or the experimental data, as applicable.**

1. The computer simulation includes pre-quantum models (Billiard Ball, Plum Pudding, etc.). It was not possible to “test” these models and actually collect emission spectra in the lab. How did the computer programmers (in consultation with expert physicists) decide on the correct output? In other words, communicate your understanding of each of the pre-quantum models and discuss how each model’s attributes led the programmers to predict their emission spectra. Information from Part A, and from your textbook or other sources, may be helpful.
2. In this experiment you calibrated a spectroscope. Suppose you analyzed the emission spectra for a second row element (Li, Be, B, etc.) in the visible light region using the same set-up. What would be the scale readings? Express your findings in a table that includes both the reported wavelengths and your converted scale readings. There are several web-sites listed emission data (see the link below).

<http://concave.stc.arizona.edu/thepoint/Interactive/spectraline.html>

1. Examine line spectra data for other elements (the web site in question 2 is useful). Do the emission spectra for the second row elements follow a periodic trend in terms of number of emission lines, or perhaps their energy?
1. Questions in **bold** should be answered directly into this packet. [↑](#footnote-ref-1)
2. A Vernier spectrometer and gas discharge lamps are used. [↑](#footnote-ref-2)