Lab 8: Let There Be Light! (**and Matter**)

**What causes color? Do my sunglasses protect my eyes from UV light? Does color from food come from the same chemical(s) as food coloring?**

**Part 1. Emission Spectroscopy: Using Light To Identify Substances**

**Prelab**
Spend 5 minutes doing the following activity. Assign a notetaker. Report to class.

1. Point the spectroscope at the lights in room. What do you see?
   Look at the spectroscope. Where is the source? What breaks up the light into its component colors? What is the detector?

2. A spectrometer measures the wavelength of light. Calibrate the spectrometer.
   a. Measure the wavelength of _____ line from ______ tube.
   b. The emission spectrum of a substance is a property of a substance. Look up the true wavelength of ______.
   c. Calculate the wavelength difference = experimental wavelength - true wavelength.
   d. Based on the wavelength difference, what is the uncertainty (± error) in the wavelength using the Vernier spectrometer?
   e. Based on this uncertainty, how many significant figures or decimal places should you report for pressure using the pressure sensor?

**Objectives**
(i) Identify a substance from an emission spectrum.

**A Brief History of Color and Atomic Structure**

In a fireworks show, you see a variety of colors from the different fireworks. It turns out that certain substances produce specific colors. For example, sodium salts produce yellow, strontium salts produce red, barium salts produce green, and copper salts produces blue. Why do these elements produce different colors?

In the 1850’s, Kirchoff and Bunsen (of Bunsen burner fame) observed that when light is shone through a heated gas emanating from one particular element of the Periodic Table, analysis of the emerging light revealed distinctive lines peculiar only to that element. They analyzed this light using a spectroscope to obtain an emission line spectrum. (Substances that emit light of all frequencies of its spectrum give a continuous spectrum.) A spectroscope contains a diffraction grating that disperses the light into its component colors. A few years later, they pointed their spectroscope in the direction of a fire in Manheim, Germany and were amazed to observe that the resulting light from the distant fire revealed the lines of barium and strontium. Note that they used a line spectrum to identify a substance. Next, they pointed their spectroscope at the sun and other stars to learn their composition. Kirchoff and Bunsen concluded that stars consist of the same elements as on earth. The key to their conclusion is the nature of color. Their work was one of the first pieces of the puzzle that led to the elucidation of the structure of the atom.

In the early 1900’s, Planck investigated black body radiation. He observed that an atom that is heated emits radiation. Based on his observations, he assumed, incredibly, a minimum quantity of energy is emitted at any given time. He called this minimum quantity, or small bundle, of energy a quantum and said that the energy of a quantum is related to the frequency of the light emitted. Einstein, from his photoelectric effect observations, showed that these bundles of light, which he called photons, travel through space and related the energy of the photon to the wavelength by which it is propagated. These observations showed that light has a dual nature - it appears to behave like a wave (it has a frequency and wavelength) or a particle (photon).

As you know, an atom consists of a nucleus and electrons. In the 1900’s, the structure of the atom was not known although there were many theories. Lord Kelvin in 1902 proposed that the structure of an atom was a homogeneous sphere of positive charge with electrons embedded in this sphere (like raisins in a cup of Jell-O). In 1910, Rutherford tested this model by directing a beam of alpha particles at a gold foil. His experiments showed most of the alpha particles passed through the gold foil and some were deflected. However, a few of the particles were deflected backwards. He concluded that the nucleus
of the atom occupies a very small volume but makes up most of its mass whereas the electrons occupy most of the volume of an atom but has a very small mass. In 1913, Bohr combined the ideas of quantization and atomic structure to develop a model for the hydrogen atom. In his model, he postulated that
1. electron have certain, discrete energies, i.e., electron energy levels are quantized, and
2. electrons can undergo transitions between energy levels.

When an atom or molecule absorbs energy, an electron undergoes a transition from a lower energy state to a higher energy state. However, since these electron energy levels are quantized, the exact amount of energy that corresponds to the energy difference between energy levels must be absorbed for this transition to occur. This is called an absorption process. An electron in a higher energy level (an excited state) does not stay there for long. It undergoes another transition from this higher energy level back to a lower energy level. In this emission process, energy is released or emitted in the form of light. The energy difference between the two energy levels determines the frequency and wavelength of the light emitted.

According to Bohr’s model, the wavelength of the lines of the H emission spectrum are described by

$$\frac{1}{\lambda} = R\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right)$$  \(1\)

where \(n_i\) and \(n_f\) are small whole numbers (\(n_f > n_i\))

\[ R = \frac{R_H}{hc} = \frac{(2.18 \times 10^{-18} \text{ J})/(6.63 \times 10^{-34} \text{ J sec})(3.00 \times 10^8 \text{ m/sec})}{1.097 \times 10^7 \text{ m}^{-1}}. \]

Since Equation (1) accurately calculated the observed H emission lines, Bohr’s model was accepted within the scientific community. The Bohr model can only be used to predict the hydrogen emission spectrum; it cannot be used to predict the emission spectrum of other atoms.

During an emission process, one energy (or frequency or wavelength) of light is emitted when an electron undergoes a transition from a higher energy level to a lower one. Since sodium salts emit yellow light, the energy difference, \(\Delta E\), between energy levels in sodium is different than in strontium, which emits red light. Think of the energy levels as steps in a flight of stairs. You must step the space between steps to move up the stairs. These energy levels, or steps, in atoms and molecules are responsible for color.

**Materials**

UV-VIS spectrophotometer

gas spectrum tubes (hydrogen, helium, argon, bromine, chlorine, iodine, mercury, neon)

light bulb

Spectroscopes and power supplies

Pickles

**Procedure**

1. Three stations will be set up.

<table>
<thead>
<tr>
<th>Station</th>
<th>Source</th>
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| 1       | a. Solid source - light bulb with W filament  
b. Unknown A |
| 2       | a. Gas source – Ne  
b. Unknown B |
| 3       | Hydrogen gas source |

At each station, use the spectroscope and the spectrophotometer connected to the computer to do the following:

a. determine whether the emission spectrum is a continuous spectrum or line spectrum.
b. For the line spectra, count the number of lines that you see.
c. Record the intensity of each line - whether it is bright (tall line) or faint (short line).

c. Record the color and wavelength of each line that you see.
Use at least three significant figures for the wavelength.

2. a. For the hydrogen emission, record your observed wavelengths and colors in Table 1 as a spreadsheet. **USE THE SPREADSHEET AS A CALCULATOR** (instead of doing the calculations on your calculator and transcribing the numbers onto a spreadsheet).

   b. Use algebra to rearrange Equation (1) so \( 1/\lambda R \) is on the left side of the equation.

For each line you observed, calculate \( 1/\lambda R \).

c. Determine the values for \( n_f \) and \( n_i \). Check your \( n_f \) and \( n_i \) values by calculating \( (1/n_f^2 - 1/n_i^2) \).

d. Calculate \( \Delta E \).

e. Draw an energy level diagram of the H atom. (The y axis is the energy scale in J.) Use an arrow to show the electronic transition from one energy state to another. Show \( \Delta E \) of each transition.

### Table 1. Hydrogen Emission Spectrum Data and Calculations

<table>
<thead>
<tr>
<th>Observed Wavelength, nm</th>
<th>Observed Color</th>
<th>( 1/\lambda R )</th>
<th>( n_f )</th>
<th>( n_i )</th>
<th>( (1/n_f^2 - 1/n_i^2) )</th>
<th>( \Delta E, J )</th>
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</table>

**Questions**

1. Show your Table 1 and your energy level diagram of the H atom.

2. a. Describe how light is produced. Draw a picture that supports your description

   b. The neon gas source showed many lines in its emission spectrum. Why does neon appear red?

   c. How is an emission spectrum of a substance like a fingerprint?

3. A sentence in the last paragraph of the Introduction states, “Think of the energy levels as steps in a flight of stairs.”

   a. How do energy levels as a flight of stairs produce a line spectrum?

   b. If a flight of stairs corresponds to a line spectrum, a ____ corresponds to a continuous spectrum.

### Problem 1

**Procedure**

1. Measure the emission spectrum of Unknown A and Unknown B.

2. Identify the element in Unknown A and Unknown B from their emission spectra.

   a. See the following references for line spectra to which you can compare each unknown.

      (i) your textbook.
      (ii) [http://physics.nist.gov/PhysRefData/ASD/index.html](http://physics.nist.gov/PhysRefData/ASD/index.html)
      (iii) [http://members.misty.com/don/spectra.html](http://members.misty.com/don/spectra.html)
      (iv) [http://www.colorado.edu/physics/2000/quantumzone/](http://www.colorado.edu/physics/2000/quantumzone/)

   b. Compare your data to the elements and wavelengths in Table 2.

3. Carefully, point the spectroscope at the fluorescent lights in the lab room. Identify the element in these lights.

4. Your instructor will demonstrate a “glowing pickle”.

   a. What color does the pickle glow?

   b. How can you use this color to identify what’s in the pickle?
Make your measurement.

c. Identify the substance that produces the color from the glowing pickle.

<table>
<thead>
<tr>
<th>Element</th>
<th>Wavelength, nm</th>
<th>Element</th>
<th>Wavelength, nm</th>
<th>Element</th>
<th>Wavelength, nm</th>
<th>Element</th>
<th>Wavelength, nm</th>
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<tr>
<td>Rn</td>
<td>745</td>
<td>Na</td>
<td>590</td>
<td>Cu</td>
<td>522</td>
<td>Zn</td>
<td>468</td>
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<tr>
<td>He</td>
<td>728</td>
<td>Na</td>
<td>589</td>
<td>Ag</td>
<td>521</td>
<td>Sr</td>
<td>461</td>
</tr>
<tr>
<td>Ar</td>
<td>707</td>
<td>He</td>
<td>588</td>
<td>Mg</td>
<td>518</td>
<td>Cs</td>
<td>459</td>
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<tr>
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<td>I</td>
<td>516</td>
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<td>456</td>
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<td>Ne</td>
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<td>Cu</td>
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<td>Ba</td>
<td>455</td>
</tr>
<tr>
<td>F</td>
<td>690</td>
<td>Hg</td>
<td>578</td>
<td>Cu</td>
<td>510</td>
<td>Zn</td>
<td>452</td>
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<td>He</td>
<td>502</td>
<td>Xe</td>
<td>450</td>
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<tr>
<td>He</td>
<td>678</td>
<td>N</td>
<td>567</td>
<td>Ba</td>
<td>493</td>
<td>Kr</td>
<td>446</td>
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<tr>
<td>Li</td>
<td>671</td>
<td>Pb</td>
<td>561</td>
<td>He</td>
<td>492</td>
<td>Kr</td>
<td>445</td>
</tr>
<tr>
<td>He</td>
<td>656</td>
<td>Kr</td>
<td>557</td>
<td>Sr</td>
<td>483</td>
<td>He</td>
<td>439</td>
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<td>644</td>
<td>Ba</td>
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<td>Hg</td>
<td>436</td>
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<tr>
<td>Ne</td>
<td>640</td>
<td>Ba</td>
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<td>Br</td>
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<tr>
<td>Li</td>
<td>610</td>
<td>Ba</td>
<td>542</td>
<td>He</td>
<td>469</td>
<td>Rb</td>
<td>420</td>
</tr>
</tbody>
</table>

**Report**

1. Report the wavelengths of each line for each Unknown.

2. Identify each Unknown.

Unknown A: ______

Unknown B: ______

Fluorescent lights: ______

Glowing pickle: ______

Show how you matched up the experimental emission lines with the emission lines of an element to identify each substance.

Would you rather use density to identify each substance or an emission spectrum?
Part 2. Absorption Spectroscopy

Introduction to Colorimetry

The relative amount of light that a substance absorbs, i.e., how light or dark the color is, is used to quantitatively measure the concentration of the color-causing substance (chromophore) in a solution. This method is known as colorimetry - the measure of color. When light of the correct wavelength (the wavelength that causes an absorption transition to occur) shines through a solution, the intensity of the light will decrease as it passes through the solution (why?). The relative intensity of light that enters and exits from the solution is defined as the transmittance, \( T \)

\[
T = \frac{I}{I_0}
\]

where \( I = \) intensity of light that exits solution
\( I_0 = \) intensity of light that enters solution.

The absorbance, \( A \), is related to the transmittance by

\[
A = -\log T = -\log \left( \frac{I}{I_0} \right)
\]

The concentration of the chromophore in solution is determined from Beer’s law. This law states that absorbance is directly proportional to the concentration of the chromophore

\[
A = kC
\]

where \( A = \) absorbance
\( k = \) proportionality constant (which depends on sample path length and the ability of the chromophore to absorb light)
\( C = \) concentration in M.

Part 2A. Using Light to Test Sunglasses and Safety Glasses

Prelab
Spend 5 minutes doing the following activity. Assign a notetaker. Report to class.
Students: bring a pair of sunglasses
An absorption spectrum measures the wavelengths and amount of light absorbed by a substance.
a. How can you use a spectroscope or spectrometer to see how well sunglasses work?
b. If your sunglasses provide “100% UV protection,” draw an absorption spectrum of these sunglasses.

Objectives
(i) Determine the UV blocking ability of sunglasses from an absorption spectrum.

Introduction
Many sunglasses claim to provide “100% UV protection” for your precious eyes. Do expensive sunglasses do a better job of blocking UV than cheap sunglasses (see ZZ Top)?

Materials
UV-VIS spectrophotometer
Students: bring a pair of sunglasses

Procedure
1. Design an experiment to measure the absorption spectrum of sunglasses and safety glasses.
2. a. Measure your absorption spectra.
b. What region of the electromagnetic spectrum is absorbed by your sunglasses?

Report the wavelength(s) of light absorbed.

c. Record your data in Table 3. Share your data with the rest of the class on the table on the chalkboard.

Table 3. Sunglasses and safety glasses absorbance data. Anything else go in this table?

<table>
<thead>
<tr>
<th>Sunglasses brand</th>
<th>Cost, $</th>
<th>Wavelength(s) of UV peaks</th>
<th>Wavelength(s) of visible peaks</th>
</tr>
</thead>
<tbody>
<tr>
<td>Safety glasses</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Questions
1. Show your Table 3.

2. Based on the results of the class, which sunglasses protect eyes the best from UV? Give reasons. Include numbers!

3. Draw a picture of how sunglasses work. Show the wavelengths of light that are absorbed by ____ and the wavelengths of light that are transmitted through the ____.
Part 2B. Identify Colored Substances from an Absorption Spectrum

Prelab
Spend 5 minutes doing the following activity. Assign a notetaker. Report to class.
1. Add one drop of food coloring in 10 ml of water and mix. In a separate container, add 1 drop in 20 ml of water and mix. In a third container, add 1 drop in 50 ml of water and mix.
   a. Which solution is the most concentrated? Which solution absorbs the most light?
   b. What color of light is absorbed by the solution? Is this color the same as the solution color?
   c. When you look at the food coloring solution, what is the light source? What is the detector?

2. Take a look at the spectrophotometer. You want to measure the absorption spectrum of food coloring in water.
   a. You measure the absorption spectrum of a “blank” solution. What is the blank solution in this example? Hint: compare this step to “zero-ing” (tare) a balance when you want to measure the mass of a substance.
   b. Then, you measure the absorption spectrum of the food coloring in water. How do you determine the absorption spectrum of the food coloring only?

   a. Do Spearmint Lifesavers, Peppermint Lifesavers, and Wintergreen Lifesavers consist of the same chemicals?
   b. If not, what chemical is in one Lifesaver but is not in another?

Figure 1. Emission spectra of Lifesavers
Figure 2. Emission spectra of Wintergreen Lifesavers (top) and methyl salicylate (bottom)

Objectives
(i) Identify colored substances from absorption spectra.
Lab 8: Let There Be Light! (and Matter)

Introduction
You looked at emission spectra in Part 1. A substance with a \( \Delta E \) between electron energy states that corresponds to a red wavelength will emit red when an electron undergoes a transition from the higher energy state to lower energy state. What happens if an electron undergoes a transition from the lower energy state to higher energy state in the same substance? A red wavelength is absorbed but the substance will not appear red. It will appear its complementary color, green. See the color wheel in Fig 1.

![Color wheel and complementary colors](http://kristiesloan.com/choosing-a-color-palette-for-a-project-complementary/)

Absorption spectra and emission spectra are like a fingerprint. You can use spectra to identify substances.

Materials
UV-VIS spectrophotometer
Students: bring a colored food or plant, e.g., colored hard candy, red cabbage, grape juice, grass, blueberries
Food coloring
test tubes
test tubes
volume measuring devices (pipet or graduated cylinders)
ethanol
hexane

Procedure
Your group will work with two different food colors. Your instructor will indicate which food coloring to use.

1. Measure the absorption spectrum of your two food colors.
   a. Prepare your food coloring solution by adding 1 drop of food color to 100 ml of water. Mix well.
      Which container should you use and how accurate should the volume be measured?
   b. Measure the absorption spectrum of this solution from 200 nm to 800 nm.
   c. From your spectrum, find the wavelength that corresponds to the absorption peak of this solution, i.e., select the wavelength at which the absorbance is greatest. Label this wavelength on your spectrum.
   d. Repeat with your second food color.
   e. Plot both food colors on the same graph.

2. a. How many lines or peaks did you observe for each food color?
   b. How does the wavelength at which you observed an absorption peak correspond to the wavelength of the food color? In other words, does red food color absorb a red wavelength? If not, explain what you observed.
   c. Draw a simple energy level diagram that represents each food color. Calculate the energy difference in J between the two energy levels.
Show this energy difference in your diagram.

d. Which food color absorbs higher energy photons?
e. Draw a picture that shows how food coloring absorbs certain wavelengths of light.
Show the wavelengths of light that are absorbed by ___ and the wavelengths of light that are transmitted through the ___.

3. You make a new food color by mixing two food colors together.
Is the new food color the result of a chemical reaction or just mixing?

a. Mix the two food colorings that you used in the first step in Part 2B by adding 1 drop of each food color to 100 ml of water.
What is the resulting color?

From this color, ___ the number of absorption peaks and the wavelength that corresponds to these peaks.

c. Measure the absorption spectrum of the resulting solution.
d. Plot your absorption spectrum of the mixture of two food colorings on the same graph as the absorption spectra of the individual food colorings. This graph should show three spectra. Label your graph and give it a title.

d. Summarize your peak wavelength data clearly in Table 4. Give Table 4 a title.


Questions
1. a. Show your Table 4 and graph of the three absorption spectra (the two individual food colors and the mixture).
b. Is the actual absorption spectrum of the mixture of two food colors the sum of the absorption spectra of the two colors that comprised the mixture or a completely different spectrum altogether? Give reasons. Include numbers!
c. Discuss whether a chemical reaction occurred to give the resulting color of the food coloring mixture. Include numbers!

2. You have a blue pen.
a. Could the blue color from this pen come from one substance? If so, you would see ____ peaks in an absorption spectrum and the peak corresponds to a ______ (what color)? wavelength.
b. Could the blue color from this pen come from two substances? If so, what would you see in an absorption spectrum of the blue pen?
c. Could the blue color from this pen come from more than two substances? If so, what would you see in an absorption spectrum of the blue pen?
Problem 2. Is the color from your food from food coloring? Extract the color from a food or plant and compare the color chemical to food colors.

Prelab
Spend 5 minutes doing the following activity. Assign a notetaker. Report to class.

**Students:** bring a colored food or plant, e.g., colored hard candy, red cabbage, grape juice, grass, blueberries

Cut a piece of your colored food or plant into small pieces and place it in a small volume of water. Stir. What happens? If nothing happens, add heat.

Cut a piece of your colored food or plant into small pieces and place it in a small volume of ethanol. Stir. What happens? If nothing happens, add heat.

Procedure
1. Design an experiment to remove (extract) the color from a food or plant.
   You can use a colored hard candy, red cabbage, grass, red grapes, etc.
   You can use water, ethanol, or hexane as the solvent.
   Clearly state the plant or food from which you extracted the color.

2. Measure the absorption spectrum of the color you removed from the food or plant.
   What is the peak wavelength?

3. Plot your absorption spectrum of the plant or food on the same graph as the absorption spectrum of the food color that most closely matches the color of the plant or food.


Questions
1. Show your absorption spectra of the plant or food and the food color. Include numbers!

2. Discuss the identity of the color from the plant or food compared to the food color. Does the color from the food or plant come from the same chemical as a food color? Give reasons. Include numbers!