Activity 14. Observing Bright Line Spectra

BACKGROUND

When atoms are excited, they absorb energy and move to a higher electron orbital. They can then give off light while dropping back into a lower electron orbital. Examination of the light given off, shows that each atomic species puts out light at distinct wavelengths, or steps. We call this stepped light quantized. Quantum theory had to examine and explain this observed property of light given off from excited atomic species.

We can excite atoms or molecules by putting the gas into a glass tube, at pressure much less than atmosphere, with electrodes at both ends. When we apply electrical current across the electrodes, the atoms are excited to higher electron orbitals. They give off light, as they return to lower orbitals. This is the basic principle behind the neon light; a noble gas is put into a low pressure tube with electrodes at each end.

If we were to take a prism or a diffraction grating and shine sunlight through it, we would see a rainbow of colors with each color merging into the next color. If we do the same thing with an incandescent light bulb, we still see the rainbow. But, if we look at the light refracted through a prism from a fluorescent light or from one of the gas tubes described above, we only see a series of lines, no rainbow. Sunlight or incandescent light show a continuous light spectra, as they are caused by the “black body effect” – a heated object gives off light, with the color being a function of the temperature. Burning coals are red, burning magnesium is white (higher temperature). A light filament or the sun gives off white light, as they are very hot. The plasma discharge tubes are not emitting light due to heating, but due to the excitation of their electrons from the electric field; therefore, we see only a few lines.

In this lab activity, we will use spectrophotometers to determine the wavelengths of light emitted from a plasma discharge tube filled with hydrogen – we will observe the hydrogen emission spectrum. We will then use the Rydberg equation to determine which electron energy transitions are being observed in the hydrogen spectrum.

A spectrophotometer (photo = light, spectro = separation, meter = measures) is used to separate the light into it’s different wavelengths so that we can analyze the light. It contains a slit, for the light to enter, a grating to separate the light and a means to measure the wavelength of the light. The unit we use requires your eyes to measure the wavelength. There is a number scale under the measurement zone. You examine where the light occurs, record the numbers and then use a known spectrum to calibrate the numbers. Other spectrophotometers use a light meter, which scans over the measurement zone (=a scanning spectrophotometer) or a photodiode array within the measurement zone. The diode array can have 512 zones, 1024 zones, or 2048 zones, or “pixels”. Again, a known spectrum is used to calibrate the system.

We use a spectrophotometer to study emission spectra (light given off by a reaction). This can be a plasma discharge, a flame discharge, light given off by distant stars, or light given off by reactions. There are a series of reference books which contain tables of spectra given off by atomic and molecular discharges. We make use of these to interpret an unknown spectrum.
Calculation of the Energy Levels of the Transitions:

In the late 1800s, Rydberg discovered that the wavelengths of the various lines in the hydrogen spectrum could be related by a mathematical equation:

\[
\frac{1}{\lambda} = R \left( \frac{1}{n_f} - \frac{1}{n_i} \right)
\]

where:  
R is Rydberg’s constant = 1.097x10^7 m\(^{-1}\)  
\(n_f\) and \(n_i\) are small whole numbers  
(the final and initial energy levels of the transition.

and \(\lambda\) is Lambda, the wavelength, in nanometers.

Note: 1.097x10^7 m\(^{-1}\) = 1.097x10^-2 nm\(^{-1}\) Use this alternative version to get nm directly.

You can also use the relationships (note that a negative sign on \(\Delta E\) will indicate the release or emission of a photon):

\[
c = \lambda \nu \quad E_{\text{photon}} = h \nu \quad E_{\text{photon}} = |\Delta E| \quad E_n = -\frac{R \hbar c}{n^2} \quad \Delta E = -R \hbar c \left( \frac{1}{n_f} - \frac{1}{n_i} \right)
\]

\[
c = 3.00 \times 10^8 \text{ m/s} \quad h = 6.626 \times 10^{-34} \text{ J s} \quad R \hbar c = 2.18 \times 10^{-18} \text{ J}
\]

Helium Spectral Lines (used to calibrate the spectrophotometer)

<table>
<thead>
<tr>
<th>Intensity</th>
<th>Wavelength (nm)</th>
</tr>
</thead>
<tbody>
<tr>
<td>500</td>
<td>388.9</td>
</tr>
<tr>
<td>200</td>
<td>447.1</td>
</tr>
<tr>
<td>500</td>
<td>587.5</td>
</tr>
<tr>
<td>100</td>
<td>667.8</td>
</tr>
<tr>
<td>200</td>
<td>706.5</td>
</tr>
</tbody>
</table>
Observing Bright Line Spectra

**Directions:** Please perform the following tasks and answer the questions in the spaces provided.

1. Calculate the energy of the following energy levels of the hydrogen atom. 
   Show work for at least 2 of the calculations.
   
   \[ n = 6 \quad \text{______________} \]  
   
   \[ n = 5 \quad \text{______________} \]  
   
   \[ n = 4 \quad \text{______________} \]  
   
   \[ n = 3 \quad \text{______________} \]  
   
   \[ n = 2 \quad \text{______________} \]  
   
   \[ n = 1 \quad \text{______________} \]  

2. Based on just these 6 different energy levels, how many different energy transitions (releasing photons) are possible? _________

3. Using the values you calculated for \( n=1 \) to \( n=6 \), calculate \( \Delta E \) for ALL of the possible transitions (from higher to lower energy states) within these six energy levels.

<table>
<thead>
<tr>
<th>Final ( n )</th>
<th>Initial ( n )</th>
<th>( 6 )</th>
<th>( 5 )</th>
<th>( 4 )</th>
<th>( 3 )</th>
<th>( 2 )</th>
</tr>
</thead>
<tbody>
<tr>
<td>5</td>
<td></td>
<td>-----</td>
<td>-----</td>
<td>-----</td>
<td>-----</td>
<td>-----</td>
</tr>
<tr>
<td>4</td>
<td></td>
<td>-----</td>
<td>-----</td>
<td>-----</td>
<td>-----</td>
<td>-----</td>
</tr>
<tr>
<td>3</td>
<td></td>
<td></td>
<td>-----</td>
<td>-----</td>
<td>-----</td>
<td></td>
</tr>
<tr>
<td>2</td>
<td></td>
<td></td>
<td></td>
<td>-----</td>
<td>-----</td>
<td></td>
</tr>
<tr>
<td>1</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td>-----</td>
<td></td>
</tr>
</tbody>
</table>
4. What is the energy of the n = ∞ energy level?

\[ n = \infty \]

5. **Observe the spectrum hydrogen from the gas discharge tubes.** Complete the following table.

<table>
<thead>
<tr>
<th>Wavelength</th>
<th>Color</th>
<th>Frequency</th>
<th>E_{ photon}</th>
<th>Transition</th>
</tr>
</thead>
<tbody>
<tr>
<td>410.2 nm</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>434.0 nm</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>486.1 nm</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>656.3 nm</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**Calculations:**

6. Why can we not observe all of the spectral lines we calculated the energies for in the Hydrogen atom?

7. **Observe the spectra of other elements.** Why do other elements have emission lines at different wavelengths from hydrogen?

<table>
<thead>
<tr>
<th>Gas</th>
<th># lines</th>
<th>color of light</th>
</tr>
</thead>
<tbody>
<tr>
<td>Helium</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Neon</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Argon</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Krypton</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Xenon</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Sodium</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>